Combined Science Trilogy 1

Combined Science Trilogy 1

Edited by James Napier

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A wide range of other useful resources can be found on the relevant subject pages of our website: www.aqa.org.uk.
## Contents

**Get the most from this book**

### Biology 1

<table>
<thead>
<tr>
<th>Chapter</th>
<th>Title</th>
<th>Page</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Cell structure</td>
<td>1</td>
</tr>
<tr>
<td>2</td>
<td>Cell division</td>
<td>18</td>
</tr>
<tr>
<td>3</td>
<td>Transport in cells</td>
<td>28</td>
</tr>
<tr>
<td>4</td>
<td>Animal tissues, organs and organ systems</td>
<td>42</td>
</tr>
<tr>
<td>5</td>
<td>Plant tissues, organs and organ systems</td>
<td>67</td>
</tr>
<tr>
<td>6</td>
<td>Infection and response</td>
<td>78</td>
</tr>
<tr>
<td>7</td>
<td>Photosynthesis</td>
<td>95</td>
</tr>
<tr>
<td>8</td>
<td>Respiration</td>
<td>104</td>
</tr>
</tbody>
</table>

### Chemistry 1

<table>
<thead>
<tr>
<th>Chapter</th>
<th>Title</th>
<th>Page</th>
</tr>
</thead>
<tbody>
<tr>
<td>9</td>
<td>Atomic structure and the periodic table</td>
<td>116</td>
</tr>
<tr>
<td>10</td>
<td>Bonding, structure and the properties of matter</td>
<td>147</td>
</tr>
<tr>
<td>11</td>
<td>Quantitative chemistry</td>
<td>176</td>
</tr>
<tr>
<td>12</td>
<td>Chemical changes</td>
<td>199</td>
</tr>
<tr>
<td>13</td>
<td>Energy changes</td>
<td>230</td>
</tr>
<tr>
<td>14</td>
<td>Formulae and equations</td>
<td>243</td>
</tr>
</tbody>
</table>

### Physics 1

<table>
<thead>
<tr>
<th>Chapter</th>
<th>Title</th>
<th>Page</th>
</tr>
</thead>
<tbody>
<tr>
<td>15</td>
<td>Energy</td>
<td>258</td>
</tr>
<tr>
<td>16</td>
<td>Electricity</td>
<td>291</td>
</tr>
<tr>
<td>17</td>
<td>Particle model of matter</td>
<td>318</td>
</tr>
<tr>
<td>18</td>
<td>Atomic structure</td>
<td>337</td>
</tr>
</tbody>
</table>

### Appendix: The periodic table

<table>
<thead>
<tr>
<th>Chapter</th>
<th>Title</th>
<th>Page</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td>357</td>
</tr>
</tbody>
</table>

### Glossary

<table>
<thead>
<tr>
<th>Chapter</th>
<th>Title</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Chapter</th>
<th>Title</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

### Index

<table>
<thead>
<tr>
<th>Chapter</th>
<th>Title</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

### Acknowledgements

<table>
<thead>
<tr>
<th>Chapter</th>
<th>Title</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Chapter</th>
<th>Title</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Chapter</th>
<th>Title</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Chapter</th>
<th>Title</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Welcome to the AQA GCSE Combined Science Trilogy Student Book 1.

This book covers the Foundation and Higher tier content for sections Biology 4.1–4.4, Chemistry 5.1–5.5 and Physics 6.1–6.4 of the 2016 AQA GCSE Combined Science Trilogy specification. All other content can be found in Book 2.

The following features have been included to help you get the most from this book.

**Prior knowledge**
This is a short list of topics you should be familiar with before starting a chapter. The questions will help to test your understanding. Extra help and practice questions can be found online in our AQA GCSE Science Teaching & Learning Resources.

**KEY TERMS**
Important words and concepts are highlighted in the text and clearly explained for you in the margin.

**Practical**
These practical-based activities will help consolidate your learning and test your practical skills.

**Required practical**
AQA’s required practicals are clearly highlighted.

**TIPS**
These highlight important facts, common misconceptions and signpost you towards other relevant chapters. They also offer useful ideas for remembering difficult topics.

**Show you can...**
Complete the Show you can tasks to prove that you are confident in your understanding of each topic.

**Higher-tier only**
Some material in this book is only required for students taking the Higher-tier examination. This content is clearly marked with the blue symbol seen here.

**Examples**
Examples of questions and calculations that feature full workings and sample answers.

**Test yourself questions**
These short questions, found throughout each chapter, allow you to check your understanding as you progress through a topic.
Chapter review questions

These questions will test your understanding of the whole chapter. They are colour coded to show the level of difficulty and also include questions to test your maths and practical skills.

- Simple questions that everyone should be able to answer without difficulty.
- These are questions that all competent students should be able to handle.
- More demanding questions for the most able students.

Practice questions

You will find Practice questions at the end of every chapter. These follow the style of some of the different types of questions you might see in your examination and have marks allocated to each question part.

Working Scientifically

In this book, Working Scientifically skills are explored in detail in the activities at the end of most chapters. Work through these activities on your own or in groups. You will develop skills such as Dealing with data, Scientific thinking and Experimental skills.

Extension

Occasionally we have included material that isn’t in the AQA specification. You can use this for further reading and deepen your understanding of a topic. This may be especially useful for students hoping to study A Level science. This content is clearly marked with the green symbol seen here.

* AQA only approve the Student Book and Student eTextbook. The other resources referenced here have not been entered in the AQA approval process.
There are thousands of different types of cell found in millions of different species of life on Earth. These range from tiny bacteria that live around us to cells in birds that can fly over the Himalayas. Cells can be put into two broad groups: prokaryotic cells found in prokaryotic organisms (also called prokaryotes) and eukaryotic cells found in eukaryotic organisms (eukaryotes). Prokaryotic and eukaryotic cells have many features in common but also some key differences.

This chapter covers specification points 4.1.1.1 to 4.1.1.5 and is called Cell structure. It covers eukaryotic and prokaryotic cells, animal and plant cells in more detail, and microscopy.
Eukaryotes and prokaryotes

**Eukaryotes**

All animal and plant cells are eukaryotic, which makes all plants and animals eukaryotes. Figure 1.1 shows examples of the huge diversity we can see in eukaryotic life on Earth.

You can see from Figure 1.1 that many eukaryotes are complex organisms. Organisms that are made from more than one cell are described as multicellular.

**Prokaryotes (bacteria)**

All bacterial cells are prokaryotic, which means that all bacteria are prokaryotes.

---

**KEY TERMS**

- **Eukaryotic cells** Cells that contain a nucleus.
- **Eukaryote** An organism that is made of eukaryotic cells (those that contain a nucleus).
- **Prokaryotic cells** Describe single-celled organisms that do not contain a nucleus.
- **Prokaryotes** Prokaryotic organisms (bacteria).

---

**Previously you could have learnt:**

- Cells are the basic unit of living organisms.
- Each part of a cell has a particular function.
- Plant and animal cells have some similarities but also some differences.
- Unicellular organisms can be adapted for particular functions.

**Test yourself on prior knowledge**

1. What are the functions of plant cell walls?
2. Describe a difference between plant and animal cells.
3. Explain why plant leaves are green.
Prokaryotes:
- are single celled
- do not have a nucleus containing their genetic material (DNA)
- are smaller than eukaryotic cells
- may also have small rings of DNA called plasmids.

Individual bacterial cells are usually between 1 µm and 10 µm in length. One million micrometres (µm) make up 1 metre (m), and 1000 make up 1 millimetre (mm). This means that between 100 and 1000 bacteria will fit in a straight line in a space of 1 mm. Groups of bacterial cells are called colonies.

\[
1 \text{ m (metre)} = 100 \text{ cm (centimetres)}
\]
\[
1 \text{ cm} = 10 \text{ mm (millimetres)}
\]
\[
1 \text{ mm} = 1000 \mu \text{m (micrometres)}
\]
\[
1 \mu \text{m} = 1000 \text{ nm (nanometres)}
\]

**Figure 1.2** A bacterial cell as seen with a microscope (magnified ×20000) and as three- and two-dimensional diagrams.

A typical bacterial cell is shown in Figure 1.2. The functions of bacterial cell components are shown in Table 1.1.
Table 1.1 The components of bacterial cells and their functions.

<table>
<thead>
<tr>
<th>Component</th>
<th>Structure and function</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cytoplasm</td>
<td>This fluid is the part of the cell inside the cell membrane. It is mainly water and it holds other components such as ribosomes. Here most of the chemical reactions in the cell happen (such as the making of proteins in ribosomes).</td>
</tr>
<tr>
<td>Cell wall</td>
<td>Like those of plants and fungi, bacterial cells have a cell wall to provide support. However, unlike plant cell walls this is not made of cellulose. The cell membrane is found on the inside surface of the cell wall.</td>
</tr>
<tr>
<td>Single DNA loop (DNA not in chromosomes)</td>
<td>DNA in prokaryotes is not arranged in complex chromosomes as in eukaryotic cells. It is not held within a nucleus.</td>
</tr>
<tr>
<td>Plasmids</td>
<td>These are small, circular sections of DNA. They have many functions, including giving bacterial cells resistance to some antibiotics.</td>
</tr>
<tr>
<td>Cell membrane</td>
<td>This controls what substances go in and out of a cell.</td>
</tr>
<tr>
<td>Ribosome</td>
<td>Proteins are made by ribosomes. Ribosomes are examples of cell organelles.</td>
</tr>
</tbody>
</table>

**KEY TERMS**

**Ribosome** A small cell organelle in the cytoplasm in which proteins are made.

**Organelle** A small structure with a specific function in a cell.

**TIP**

It is important that you can explain what a chromosome is.

**Figure 1.3** Prokaryotic bacterial cells seen on a pinhead.

**Test yourself**

1. Describe how DNA is arranged in bacteria.
2. Describe the function of ribosomes.

**Figure 1.3** shows how small bacterial cells are. Typical eukaryotic cells are much larger than this. However, even eukaryotic cells are microscopic. This means you can’t see a single cell without using a microscope.

**Animal and plant cells**

**Generalised (typical) animal cells**

Plant and animal cells are eukaryotic. Eukaryotic cells almost always have a nucleus and are generally larger than prokaryotic cells.

The structure of a generalised animal cell is shown in Figure 1.4.
Components of animal cells

In the previous section we looked at bacterial cells. Animal cells, including human cells, have many components in common with these. The cytoplasm of animal cells is also mainly water and it holds other components such as ribosomes. In the cytoplasm most of the chemical reactions in the cell happen (such as the making of proteins using ribosomes).

The cell membrane of animal cells also surrounds the cell. There are no cell walls in animal cells and so the membrane is on the outside of these cells. The membrane controls what substances go in and out of the cell. Most cells need glucose and oxygen for respiration, and these substances move by diffusion or are transported into the cells from the blood, where they are found at a higher concentration. Carbon dioxide moves back into the blood capillaries through the membrane.

Mitochondria are small organelles found in the cytoplasm and are only present in eukaryotic cells. They are the site of aerobic respiration. Here the energy stored in glucose is released, using oxygen. More active cells, such as those in muscles or sperm cells, usually have more mitochondria because these cells need more energy.

Ribosomes are the site of protein synthesis. These organelles are present in the cytoplasm of animal cells.

Animal cells are unlike bacterial cells in that they usually possess a nucleus. This component is present in almost all eukaryotic cells. It is found in the cytoplasm and is surrounded by its own membrane. The cell’s genetic material (DNA) is enclosed within it, arranged into chromosomes. The nucleus controls the activities of the cell.
**Generalised (typical) plant cells**

Like animal cells, plant cells are eukaryotic. They have a nucleus and they are generally larger than prokaryotic (bacterial) cells.

**Components of plant cells**

Plant cells have many components in common with animal cells, including a **nucleus** in which the organism’s genetic material (DNA) is found. As in animal cells, the DNA is packaged into chromosomes. Plant cells also have **ribosomes** for protein synthesis and **mitochondria** for respiration, in their cytoplasm.

Plant cells have some components not present in animal cells. **Chloroplasts** are small organelles, full of a green pigment called **chlorophyll**, which absorb the light necessary for **photosynthesis** to occur. This reaction uses the light energy from the Sun to convert carbon dioxide and water into glucose and oxygen and so provides an energy source for the plant. It is the green chlorophyll in plants that gives some of their parts their green colour. Most roots are hidden from the Sun and so cannot photosynthesise. They do not have chloroplasts and so are often white, not green.

Plant cells also have a **cell wall**, unlike animal cells. This is made from **cellulose** and provides structure for the cell. Plants would not be able to stand upright to catch light energy from the Sun without cell walls. The **cell membrane** is found inside the cell wall.

Many plant cells also contain a **permanent vacuole**. This is filled with cell sap (water in which dissolved sugars and mineral ions are found). The pressure in the vacuole presses the cytoplasm against the wall to keep the cell **turgid**.

**KEY TERMS**

- **Photosynthesis** A chemical reaction that occurs in the chloroplasts of plants and algae and stores energy as glucose or starch.
- **Turgid** Describes cells that have a lot of water in their vacuole. The pressure created on the cell wall keeps the cell rigid.

**TIP**

Make a model of a bacterial, plant or animal cell, and label it with the cell components and their functions to help you remember them.

Most plant cells have chloroplasts, a permanent vacuole and a cell wall made of cellulose; these features are **not** present in animal cells.
Use a light microscope to observe, draw and label a selection of plant and animal cells

In this practical you will examine the structure and features of different animal and plant cells.

Your teacher may provide you with slides showing a range of cells from plants and animals. If this is the case, use Method 1 below.

**Method 1**

1. Place your slide on a microscope stage and observe using the lowest power objective lens.
2. Focus in on the image and then increase the magnification until you can clearly observe the cell’s structure.
3. Make a drawing of what you observe, labelling any structures you recognise. Ensure that you record the magnification you used when making your observations.

Alternatively, your teacher may ask you to make up your own slides to examine the cells in a range of tissues. If this is the case, use Method 2 below.

**Method 2**

**Examining plant cells**

1. Wear eye protection.
2. Use tweezers to remove a thin sheet of cells (epidermal tissue) from the inner part of an onion layer.
3. Place this flat on a microscope slide, being careful not to fold it.
4. Place a drop of iodine onto the onion tissue.
5. Carefully lower a cover slip on top of the tissue, ensuring no air bubbles form (Figure 1.7).
6. Follow the steps in Method 1 above to examine and draw the cells present.

7. Repeat this process using a leaf from a piece of pondweed (*Elodea*), but add a drop of water rather than iodine.

**Examining animal cells**

1. Wear eye protection.
2. Using an interdental stick or flossing brush from a freshly opened pack, gently scrape the inside of your cheek.
3. Smear the cotton swab on the centre of the microscope slide in small circles.
Cell specialisation

The previous section looked at generalised animal and plant cells. Eukaryotic organisms like us are not usually made only of generalised cells. We have developed specialised cells that have adaptations to allow them to complete specific functions. Red blood cells, for example, have a biconcave shape (which dips in the middle on both sides) to allow oxygen to be absorbed more quickly. They also have no nucleus, which means they can absorb more oxygen. Some specialised cells in animals and plants, together with their adaptations, are listed below.

**Sperm cell**

In humans, about a teaspoon of semen is ejaculated during a male orgasm. In the semen are tens of millions of sperm cells, which must swim through the female reproductive system. Here one cell may fertilise an ovum (egg cell). Sperm cells have a tail to help them swim towards the ovum (Figure 1.8). They have a relatively large number of mitochondria to release the energy needed to help them swim. The nucleus of a human sperm contains the genetic material (DNA) of the father.

**Test yourself**

3 Name two structures present only in plant cells.
4 In which types of cell would you find mitochondria?
5 Describe the function of the cytoplasm.

**Show you can...**

Describe the function of mitochondria.

**Key terms**

Biconcave Describes a shape with a dip that curves inwards on both sides.

Ova (singular ovum) Eggs.

**Tip**

If you are asked to explain how a cell is adapted to its function, don’t forget to use a connecting phrase like ‘so that it can’ or ‘to allow it to’.

**Figure 1.8 The parts of a sperm cell.**
Nerve cell

Our nervous system controls and coordinates all our actions. Nerve impulses are electrical signals that travel along nerve cells. Some of our nerve cells are the longest cells in our body. Their long extensions are called axons and these have a myelin sheath surrounding much of their length (Figure 1.9). This acts like the plastic coating on an electrical wire and insulates the electrical impulse. The cell body of the nerve cell also has smaller extensions which allow it to pick up signals from neighbouring cells.

Figure 1.9 The parts of a nerve cell.

Muscle cell

Muscle cells are specialised cells that can contract and so move parts of the body. Muscle cells contain large numbers of mitochondria, as muscular contraction requires a lot of energy.

Figure 1.10 A model of nerve cells made from sweets.

Figure 1.10 A model of nerve cells made from sweets.

TIP

When muscle cells contract they get smaller.

TIP

The reason why we can make large movements, such as bending our arms, is that many muscle cells contract at once in a muscle.

Figure 1.11 The parts of a root hair cell.

Root hair cell

Root hair cells in plants have a small thin extension which pokes out into the soil (Figure 1.11). The purpose of these hairs is to increase the surface area of the root that is in contact with the soil. This allows the plant to absorb more water and minerals from the soil. Without root hairs it is likely that the adult plant would not be able to absorb enough water to survive.

Figure 1.11 The parts of a root hair cell.
Xylem cells form long tubes running along the roots and stems of plants. They carry water and some dissolved minerals from the roots upwards to other parts of the plant. This water evaporates and is lost from leaves as water vapour during the continual process of transpiration. Xylem cells also carry water to the green parts of plants for photosynthesis during the day. Xylem tubes are made from lots of individual cells that have died and have no end walls and no contents, leaving a hollow tube like a pipe (Figure 1.12). They have reinforced side walls to support the weight of the plant. The side walls are strengthened by a substance called lignin.

Phloem cells carry the glucose (as sucrose) made in photosynthesis from the leaves of a plant to all other parts of the plant in cell sap. This process is called translocation. Unlike xylem, phloem cells are living. They have fewer cell organelles than many other types of cell, which allows the sugar to travel easily. Rather than having no end walls (as in xylem), phloem cells have specialised end walls called sieve plates that have small holes in them (Figure 1.13).

Phloem cells are located close to xylem cells in the plant, making up the veins you can see in a leaf (Figure 1.14).

Cell differentiation

The previous two sections have looked at generalised and specialised animal and plant cells. After generalised cells are formed, many become specialised for specific functions. This process is called cell differentiation. Your cells did this while you were in your mother’s uterus. Part of cell differentiation involves cells developing specific structures within them to allow them to function. For example,
muscle cells need to release lots of energy during respiration and so require a high number of mitochondria. Sperm cells and nerve cells are very specialised cells. Most types of animal cell differentiate very early in their development, but most plant cells can differentiate at any stage. This is why it is possible to take plant cuttings from different parts of a plant.

**Test yourself**

6 Give the function of nerve cells.
7 Describe how nerve cells are adapted for their function.
8 Describe how root hair cells are adapted to their function.

**Microscopy**

Light microscopes use objective and eyepiece lenses to magnify structures that allow light to pass through them. The light rays travel up through the specimen and are then magnified by the objective and eyepiece lenses.

**Light microscopes**

The parts of a light microscope and their functions are shown in Table 1.2.

<table>
<thead>
<tr>
<th>Part</th>
<th>Function</th>
</tr>
</thead>
<tbody>
<tr>
<td>Eyepiece lens</td>
<td>You look through this lens to see your sample. This is often ×10.</td>
</tr>
<tr>
<td>Objective lens</td>
<td>Usually there are three to choose from (often ×5, ×10 and ×25). The smallest will be the easiest to focus, so select this first. When you have focused this lens try a different one with a greater magnification.</td>
</tr>
<tr>
<td>Stage</td>
<td>This holds the sample securely, often using two metal clips.</td>
</tr>
<tr>
<td>Specimen</td>
<td>This is usually placed in a drop of water or stain on a microscope slide under a very thin glass cover slip.</td>
</tr>
<tr>
<td>Mirror or light source</td>
<td>This sends light up through the specimen and then through the objective and eyepiece lenses into your eyes.</td>
</tr>
<tr>
<td>Coarse focus</td>
<td>This quickly and easily moves the stage up and down to focus on the sample.</td>
</tr>
<tr>
<td>Fine focus</td>
<td>This sensitively and slowly moves the stage up and down to allow you to make your image very sharp.</td>
</tr>
</tbody>
</table>

The total magnification of the image you are looking at is calculated by:

\[
\text{total magnification} = \frac{\text{magnification of eyepiece lens}}{\text{magnification of objective lens}}
\]
Electron microscopes use electrons in place of rays of light to make an image (Figure 1.16). The wavelength of electrons can be up to 100,000 times smaller than that of visible light. This means that electron microscopes can take images at significantly higher magnifications (Figure 1.17).

Figure 1.16 A scientist using a large modern electron microscope.

Figure 1.17 Images taken using an electron microscope: (a) a three-dimensional image of an ant’s head; (b) a two-dimensional image of a mitochondrion. These have had colour added.

Electron microscopes can magnify much more than light microscopes, but the key thing is that they have a much greater resolution. The resolution of microscopes is the shortest distance between two parts of a specimen that can be seen as two distinctly separate points. An electron microscope can resolve points up to 2000 times closer than a light microscope, at a separation of just 0.1 nm. The greater resolution of an electron microscope means that sub-cellular structures can be seen in much finer detail. Electron microscopes were very important in enabling scientists to understand many sub-cellular structures. For example, many organelles cannot be seen in detail with a light microscope.

Test yourself

9 State how the resolution of light and electron microscopes differs.
Chapter review questions

1. Name two features of prokaryotic cells.
2. Name the substances present in a plant cell vacuole.
3. Explain why plant cells are often green.
4. Describe how the structure of sperm cells helps their function.
5. Describe how you would make a microscope slide to look at an onion cell.

6. Describe three differences between prokaryotic and eukaryotic cells.
7. Give the function of cytoplasm and what it is made from.
8. What is the function of ribosomes?
9. Describe three differences between plant and animal cells.
10. Describe how a nerve cell is adapted for its function.
11. Define the term ‘resolution’.
12. a) Use the scale in Figure 1.18 to estimate the length of the sperm cell.
   b) Explain why the sperm cell will have a large number of mitochondria.
13. Explain why some cells have more mitochondria than other cells.
14. Define the term ‘turgid’.
15. Explain how xylem cells are adapted for their function.
16. a) The root hair cell (X) in Figure 1.19 is 1.3 mm long. By how much has it been magnified?
   b) Explain how this cell is adapted.
17. Suggest why ribosomes are usually measured in nanometres (nm).
1 Cell structure

Practice questions

1 Orchids are often found growing high up on other plants. They are unusual plants in that some species have green roots.

a) Choose the name of the structures that make parts of plants green:
   A chloroplasts
   B chlorophyll
   C mitochondria
   D ribosomes

b) Suggest why orchids grow on other plants. 

c) Suggest why some orchids have green roots.

d) Explain how root hair cells are adapted for their function.

2 Life exists on Earth as single-celled or multicellular organisms. Bacteria are single-celled organisms that grow in many places.

a) Copy the diagram of a bacterial cell in Figure 1.20, and complete the missing labels.

b) Which two of the following cell components are not found in prokaryotic organisms?
   A cell wall
   B DNA
   C nucleus
   D mitochondria

3 Microscopes have been around since the end of the 16th century. Their invention allowed us to see single-celled organisms for the first time and also understand that multicellular organisms are made up of many cells.

a) Identify parts A, B and C in the diagram of a light microscope in Figure 1.21.

b) Choose the part of the microscope that light first passes through:
   A fine focus
   B objective lens
   C eyepiece lens
   D slide

3) How is the total magnification of a light microscope calculated?

d) Describe two differences between a light microscope and an electron microscope.

4 Describe the similarities and differences between prokaryotic cells and eukaryotic plant and animal cells.
Microscopy and magnification
It is important that you can carry out calculations involving magnifications, real size and image size.

Magnification is a measure of how many times an object has been enlarged. If a sesame seed is actually 3 mm long, but in a diagram has been drawn to be 3 cm long, then it has been magnified 10 times. You can work out magnification using the formula:

\[
magnification = \frac{\text{image size}}{\text{actual size}}
\]

For example, this drawing of a flea is 40 mm long but the actual flea is 2 mm.

\(\text{Figure 1.22}\)

To work out the magnification the above formula is used:

\[
magnification = \frac{40 \text{ mm}}{2 \text{ mm}} = \times 20
\]

Sometimes you might want to know the actual size of an object if you know the magnification and size of the image. To work this out the formula for magnification can be rearranged:

\[
\text{actual size} = \frac{\text{image size}}{\text{magnification}}
\]

Also, you might need to work out what image size would be produced if you were given the actual size of the image and its magnification:

\[
\text{image size} = \text{actual size} \times \text{magnification}
\]

A formula triangle can be used to help you rearrange the equation.

\(\text{TIP}\)
It is really important to ensure that measurements are always in the same units. So if you have mixed units you will need to convert them all to the same format.

\(\text{Figure 1.23 A formula triangle.}\)
Often the actual object being studied is too small to be measured using a ruler, which means that a scale lower than a millimetre is needed. A micrometre (µm) is a thousandth of a millimetre and a millionth of a metre. Using standard form, this can be written as:

\[ 1 \, \mu m = 1 \times 10^{-3} \, \text{mm} \]
\[ 1 \, \mu m = 1 \times 10^{-6} \, \text{m} \]

**Questions**

1. If a pinhead measures 1.8 mm and is magnified \( \times 12.5 \), how large would the image be?

2. If an image of a snake’s fang is 22.5 cm and it has been magnified \( \times 7.5 \), how large is the actual fang?

3. What is the actual size of this frog’s eye if the image has been magnified \( \times 1.5 \)?

**Extension**

Cell lengths are usually measured in µm (micrometres). Sub-cellular structures can be measured in mm or nm (nanometres), depending on their size.

**Figure 1.24**

**Figure 1.25**
Example
If the actual size of this cheek cell is 60 µm, by how much has it been magnified?
- First measure the size of the cell in mm. In this micrograph, the cell is 45 mm wide.
- Then convert this to µm by multiplying by 1000.
  So 45 × 1000 = 45 000 µm
- To work out the magnification:
  \[ \text{magnification} = \frac{\text{image size}}{\text{real size}} \]
  \[ \text{magnification} = \frac{45 000}{60} \]
  \[ \text{magnification} = \times 750 \]

Question
4. What is the actual size of this red blood cell if it has been magnified ×6000?

Answer
- First measure the size of the cell in mm. In this micrograph, the cell is 45 mm wide.
- Then convert this to µm by multiplying by 1000.
  So 45 × 1000 = 45 000 µm
- To work out the magnification:
  \[ \text{magnification} = \frac{\text{image size}}{\text{real size}} \]
  \[ \text{magnification} = \frac{45 000}{60} \]
  \[ \text{magnification} = \times 750 \]
When your father’s sperm fused with your mother’s ovum you were only a single stem cell. Years later, you are now made from thousands of billions of cells. These are arranged in a very specific way and specialised into several hundred different types. This chapter explains how your body cells grew from that single fertilised ovum by a process of cell division called mitosis and how the same process replaces your damaged tissues and those cells that die naturally. This process is key to life on Earth.

This chapter covers specification points 4.1.2.1 to 4.1.2.3 and is called Cell division.

It covers the structure of chromosomes, mitosis, stem cells and cell differentiation.
Eukaryotic cells are those that contain a nucleus. These can either be single-celled organisms such as protozoa, or single cells of larger multicellular organisms such as trees, insects and you.

Almost all of the cells in your body have a nucleus in which your genetic material (DNA) is found. Half of this came from your father carried by his sperm, and the other half came from your mother in her ovum. Sex cells (sperm and ova) are called gametes, and these have only half of an organism’s DNA in them. They are described as haploid cells. Apart from gametes and some cells such as red blood cells that have no nucleus, most of your cells contain two sets of DNA: half from your mother and half from your father. Any cell with these two copies is described as a diploid cell.

The DNA in any one of your body cells (not the gametes) stretches to approximately 2 metres long. Almost all the cells in your body are so small they can only be seen using a microscope. In order to fit all of this DNA into a cell this small, it is coiled into structures called chromosomes. Humans have 23 pairs of chromosomes. We say chromosomes come in pairs to remind ourselves that half were inherited from each parent. This means that there are 46 chromosomes in a diploid human cell. This is called the ‘chromosome number’. Other eukaryotic animals and plants have different chromosome numbers. For example, mosquitos have a chromosome number of 2. This means they have one pair of chromosomes, one from each parent.

The 23 pairs of chromosomes present in a human body cell are numbered in order of their size (Figure 2.1), which varies considerably. Each of these chromosomes is divided into separate regions called genes. Each gene contains the genetic instructions to make a protein and therefore produce a characteristic. Because you have two copies of each chromosome, you also have two copies of each gene: one from your mother and one from your father.
It is important that you can explain what a gene is.

**Show you can...**

Explain why we often say that we have 23 pairs of chromosomes, rather than 46 chromosomes.

**Test yourself**

1. What is the chromosome number of humans?
2. Define the term ‘haploid’.
3. Define the term ‘diploid’.
4. Describe the difference between a gene and a chromosome.

---

**Mitosis and the cell cycle**

House dust is mainly made from our dead skin cells, and those of our family or pets who live with us. Our houses are always becoming dusty, which means that our skin is continually dying and falling off. But we don't run out of skin. This means that we must be continually replacing our dead skin cells as they fall off. This replacement process is carried out by a type of cell division called **mitosis**. This process copies a diploid body cell, which contains all of your DNA, giving two new cells identical to the original cell and each other. Without this process we would not be able to **grow** from a fertilised ovum, **repair** ourselves from damage, or **replace** the cells that die naturally throughout our lives.

**Steps in mitosis**

Figure 2.2 shows the steps in mitosis. At the top of the diagram, you can see one cell with four chromosomes. The two red ones come from one parent and the two blue ones come from the second parent. The small blue and red chromosomes make one **pair**, and the large blue and red chromosomes make the other **pair**. If this were a human cell it would have 23 pairs, but this would be too confusing in a diagram.

The first step in mitosis is for the membrane around the nucleus to disappear and all the chromosomes to shorten and fatten. This helps make the following steps easier. The chromosomes have already copied themselves completely. At this point the cells contain 46 chromosomes and 46 copies. You can see in the second box in Figure 2.2 that each of the four chromosomes now looks like an X-shape. Each of these is a chromosome with its copy.
The chromosomes and their copies then migrate to the middle of each cell, which is shown in stage 2 in the diagram. The chromosomes and their copies split apart. The chromosomes are pulled to one side of the cell and the copies to the other. This is shown in stage 3 in the diagram. The cell membrane then starts to pinch inwards and eventually touches the other side, and splits into two identical cells. We call these daughter cells. Each new daughter cell is an exact copy of the original cell.

Before a cell can undergo mitosis, it needs to:

- grow and increase the number of sub-cellular structures such as ribosomes and mitochondria so that each daughter cell gets enough
- replicate (double) the amount of DNA through the duplication of each chromosome.

Mitosis is only one part of the sequence of cell growth, increase in the number of sub-cellular structures, duplication of DNA, and cell division. The whole sequence is referred to as the cell cycle.

Cell division by mitosis is important in the growth and development of multicellular organisms. Mitosis takes place where:

- new cells are being formed during growth
- parts of the body are being repaired or replaced.

▲ Figure 2.3 The steps in mitosis as seen in bluebell plant cells.
Stem cells

**KEY TERMS**

- **Stem cell** An undifferentiated cell that can develop into one or more types of specialised cell.
- **Differentiate** To specialise, or adapt for a particular function.

**TIP**

It is important that you are able to explain the importance of cell differentiation for the examples given in this chapter.

__Stem cells in mammals__

A **stem cell** is a cell that can **differentiate** into any other type of cell. In mammals there are two types of stem cell. The first type are **embryonic stem cells**, and these cells were present when you were a zygote and an embryo. They divide rapidly by mitosis and begin to differentiate within several days of the sperm fertilising the ovum. Within 21 to 22 days the human embryo has enough differentiated cells to form a beating heart.

Embryonic stem cells can grow into **most of the specialised** cells found in the adult organism. Once an embryonic stem cell has differentiated into a specialised cell it cannot change back or turn into any other type of cell.

We have a second type of stem cells, which are simply called stem cells or **adult stem cells** (although they are also found in children). These stem cells grow only in specific parts of the body, such as the **bone marrow**. They are used to repair the body when it is injured. Crucially they develop into the types of cell found in that location. So adult blood stem cells can only develop into red or white blood cells and some other types of cell. They cannot turn into any cell in the way that embryonic stem cells can. Scientists find these cells interesting to study, but they may not be as potentially useful as embryonic stem cells.

Some animals have stem cells that allow them to regenerate parts of their body. For example, lizards can shed, and later regrow, their tail if seized by a predator. Other animals can go even further than this. If one leg of a starfish is severed by a predator it will grow back.
Stem cells

● Stem cells and differentiation in plants

Plants also have stem cells. However, those found in plants keep their ability to specialise into any type of cell. In a plant, the stem cells are located in a region called the **meristem**. This is where much of the plant’s growth occurs. Meristems are found in shoot tips, where they encourage the shoots to grow towards the light. They are also found in the root tips, where they encourage the roots to grow downwards towards water.

The fact that plant stem cells can differentiate into other cells throughout the mature organism’s life allows us to take cuttings. Here a small section of stem, usually with a few leaves, is removed. This is often dipped into rooting powder (Figure 2.6), which contains plant **hormones** to speed up differentiation. This cutting is then placed directly into the soil. The stem cells towards the bottom of the cutting will quickly grow into root cells and grow downwards. A little later we have a genetically identical copy of the parent plant, often described as a **clone**. Although there is no **genetic variation**, because of **environmental variation** the clone will not always look identical to the parent organism. Much of our food that comes from plants is grown from clones following this method. Many rare or valuable plants such as orchids and roses are grown in this way to stop them becoming too rare or becoming **extinct**. Also, stem cells can produce plants that all have good characteristics such as **disease resistance**.

○ Stem cell research

Stem cell research is an **ethical issue**. This means that some people disagree with it for religious or moral reasons. Many scientists think that research into the medical uses of stem cells might help:

- treat **paralysed patients** by making new nerve cells to transplant into a severed spinal cord or damaged brain
- treat conditions such as **diabetes** by replacing the cells in the body that are no longer working properly
- replace injured or defective **organs**.

**KEY TERMS**

**Meristem** An area of a plant in which rapid cell division occurs, normally in the tip of a root or shoot.

**Hormone** A chemical (produced in a gland in mammals) that moves around an organism to change the function of target cells, tissues or organs.

**Clone** An organism produced asexually that has identical genes to its parent.

**Genetic variation** Inherited differences in organisms.

**Environmental variation** Differences in organisms as a result of the environment in which they live.

**Extinct** When no members of a species remain alive.

**Ethical issue** An idea some people disagree with for religious or moral reasons.

**TIP**

Plant cloning can produce plants with **good characteristics** – quickly and economically.

▲ Figure 2.5 Starfish can regenerate one or more legs they have lost.

▲ Figure 2.6 A plant cutting being dipped into powder containing hormones to help the clone develop roots before being gently placed into compost.

▲ Figure 2.6 A plant cutting being dipped into powder containing hormones to help the clone develop roots before being gently placed into compost.
By using stem cells from an injured person’s own body to repair damaged tissue in a process called therapeutic cloning, doctors can be sure that the cells will not be rejected in the way that some transplants are. In the future, treatments involving therapeutic cloning of stem cells might be used to treat many medical conditions.

The most useful stem cells for this research are embryonic, because they can develop into any type of cell. These are often collected from waste cells in the left-over umbilical cord after a mother has given birth. They are found in fertilised ova that are not selected to be put into a woman’s uterus during in vitro fertilisation (IVF). It is with this that some people have an ethical issue. Some believe that a fertilised ovum is a life. Some believe that a fertilised ovum has rights and that its use in medical research amounts to murder. For this reason the regulations surrounding stem cell research are extremely tight, and some countries forbid it completely. Using stem cells can have other problems too, such as causing viral infections when infected stem cells are used.

**Key Term**

**In vitro fertilisation (IVF)** A medical procedure in which ova are fertilised outside of a woman, then placed into her uterus to develop into a baby.

**TIP**

It is important that you can discuss the benefits and risks of the use of stem cells. You should also be able to discuss the social and ethical issues regarding their use.

**TIP**

The babies born from IVF are sometimes called test-tube babies, even though no test tubes are used in the process.

**Activity**

Stem cell research is one of the most hotly debated areas of medical science. Since the first isolation of embryonic stem cells from mice in the 1980s, there has been great advances in understanding of stem cells and their potential uses in medicine. Alongside this there has been much controversy about the ethical implications of stem cell use, particularly stem cells obtained from embryos.

The UK Government is funding stem cell research and wants the UK to be a world leader in such research.

**Questions**

1 Working in small groups, come up with a list of views that different members of society might have for and against stem cell research.
2 Write each reason on a sticky note or piece of paper and rank the reasons based on which are the strongest and weakest arguments for and against stem cell research.
3 Write a letter to the Government expressing your views about stem cell research and whether you feel the Government should be funding it. Ensure you support your views with reasons. You may like to use the internet to find extra information from a range of sources to support your arguments.

**Show you can...**

Explain why stem cell research is of benefit.

**Test yourself**

7 Define the term ‘stem cell’.
8 Name the two types of stem cell.
9 Describe why stem cell research is an ethical issue.
10 Describe the advantages of using embryonic stem cells in research.
Chapter review questions

1. Give the collective term for sperm and ova.
2. What is the number of chromosomes in a human skin cell?
3. How many cells are at the beginning and end of one mitotic division?
4. Give two examples of specialised cells in animals.
5. Define the term ‘clone’.
6. Explain why plant cuttings are clones.
7. Give two examples of diploid human cells.

8. How many chromosomes are there in a human gamete?
9. Give the common name for ova.
10. Define the term ‘gene’.
11. Give two purposes of mitosis.

12. Explain why gametes need to be haploid.
13. Define the term ‘cell differentiation’.
14. Name the region of a plant in which cell differentiation occurs.
15. Suggest why some people protest against stem cell research.
Practice questions

1. How many chromosomes are there in a human diploid cell? [1 mark]
   A 48  
   B 46  
   C 44  
   D 42

2. Which of these cells are haploid? [1 mark]
   A Nerve cell  
   B Epithelial cell  
   C Sperm cell  
   D White blood cell

3. Figure 2.9 shows the nucleus of a cell that is starting to divide.
   a) Name structure A. [1 mark]
   b) Draw a diagram to show the appearance of a nucleus from a cell produced by mitosis of the cell in Figure 2.9. [2 marks]

4. Figure 2.10 shows a section through human skin.
   a) Name the type of cell division that produces new skin cells. [1 mark]
   b) Give a reason why it is important for skin cells to be able to divide. [1 mark]
   c) Describe what must happen to the genetic material before a skin cell can divide. [1 mark]

5. Scientists believe that stem cells could have many potential uses in medicine.
   a) Stem cells are described as being undifferentiated cells. What does this mean? [1 mark]
   b) Stem cells found in liver tissue are called adult stem cells. These cells are often used to repair the body. Suggest another source of adult stem cells other than the liver. [1 mark]
   c) Embryonic stem cells are useful in medicine. Give a reason why. [1 mark]
   d) Many people have differing ethical views on the use of embryonic stem cells. Suggest two reasons why some people are against the use of these cells. [2 marks]
Hypotheses and predictions

Humans have a total of 46 chromosomes, arranged in 23 pairs. Not all living things have this many: some have more chromosomes and some have fewer. Since the 1900s, when scientists were beginning to observe chromosomes more closely, they have hypothesised and made predictions linked to chromosome number.

Questions

1. Using the data in Table 2.1, come up with a hypothesis that explains the reason for the chromosome number in the animals given.

2. Scientists use their hypotheses to make predictions. Predict the diploid number of chromosomes in an elephant.

3. Predict the diploid number of chromosomes in a hedgehog.

4. Predict the diploid number of chromosomes in a goldfish.

Your hypothesis is probably linked to the idea that more complex organisms will have more chromosomes and maybe even that because they are larger they need more DNA (genes) to code for the greater amount of proteins they need to produce.

This is exactly what scientists originally thought, and indeed if you predicted that an elephant would have more chromosomes than a human you would be right: it has 56 diploid chromosomes. However, your hypothesis and prediction would not be supported by the data for the hedgehog, which has a total of 90 chromosomes, or a goldfish, which has 94.

As more and more organisms have had their chromosome numbers determined it has become apparent that there is no link between the complexity of organisms and chromosome number. This means that scientists have rejected their original hypothesis and are trying to come up with new ideas to explain the variations seen.

Today, scientists still don’t know the exact reason for differing chromosome numbers in organisms, but it is hypothesised that it is linked to their evolution and mutations that occurred in common ancestors.

For example, in some organisms two chromosomes can become fused together. This fusion of chromosomes is thought to explain the differences between chromosome number in humans and great apes. It is hypothesised that the ancestor of humans and apes had 48 chromosomes in 24 pairs but that in humans two of the chromosomes became fused so that we ended up with 46 chromosomes in 23 pairs, while the great apes still have 48 chromosomes in 24 pairs.

Scientists made a prediction that if two chromosomes had fused to make one, we should see similarities in the gene banding on the one human chromosome compared with that found on the two separate chromosomes.

Question

5. What other evidence from Figure 2.11 supports this hypothesis of two chromosomes fusing?
What does smelling your best friend’s deodorant have in common with making a cup of tea? One thing is that they both involve the movement of particles. Particles spread out naturally from areas of high concentration to areas of low concentration. This movement is called diffusion and it is a key biological process. Without it your cells would not receive oxygen or glucose, and would quickly die.

This chapter covers specification points 4.1.3.1 to 4.1.3.3 and is called Transport in cells. It covers diffusion, osmosis and active transport.
Diffusion

Diffusion is the process by which particles of gases or liquids spread out from an area where there are lots of them to areas where there are fewer of them. We say that areas with lots of particles have a high concentration and areas with fewer particles have a low concentration. This process happens naturally, and no additional energy is needed for it to occur. It is a **passive process**. Diffusion is defined as the **net** movement of particles from an area of high concentration to an area of lower concentration. Because the movement is from high to low concentration, we say it is down a **concentration gradient**.

### Examples of diffusion

When you put your deodorant on in the morning the highest concentration is under your arm. But not all of the deodorant particles remain under your arm, or nobody would be able to smell them. They slowly move by diffusion (we say they diffuse) from a high concentration under your arm to the lower concentration found in the air. The same is true of tea particles when you add hot water to your teabag. The particles of tea don’t remain in the high concentration within the bag – they spread out to the lower concentration found in the boiling water.

In the body, **diffusion occurs across cell membranes**. A good place to study diffusion is in the lungs.

---

**KEY TERMS**

- **Net** Overall.
- **Concentration gradient** A measurement of how a concentration of a substance changes from one place to another.

**TIP**

We say ‘net’ movement, which means overall movement, because some of the particles may naturally diffuse back to the area of high concentration they have just come from. Far fewer will ever do this than diffuse away, however.

---

**Previously you could have learnt:**

- Diffusion can be explained in terms of the particle model.
- Breathing is the movement of air in and out of the lungs.
- Plants gain mineral nutrients and water from the soil via their roots.

**Test yourself on prior knowledge**

1. Name the specialised cell that helps plants absorb water.
2. Describe the process of diffusion.
3. Explain why breathing and respiration are not the same thing.

---

**Figure 3.1** Molecules in a gas spread out by diffusion.

**Figure 3.2** An everyday example of diffusion. The particles move from a high concentration in a scented candle to a low concentration in the air.
Diffusion in the lungs

When we breathe in we take air that is relatively high in oxygen into our lungs. In the individual alveoli in our lungs the oxygen is in a higher concentration than in the blood, so the oxygen naturally diffuses from inside the alveoli into the blood. Because the blood in our body is always moving, the blood that now has a high concentration of oxygen is immediately moved away to the tissues and organs and is replaced by ‘new’ blood with lower levels of oxygen. This means more oxygen will always diffuse from the alveoli into the ‘new’ blood, keeping the blood rich in oxygen.

The reverse is also true when the blood that is now high in oxygen reaches our tissues and organs. Here it travels through the tiny capillaries that supply the tissues and cells. These cells have a low concentration of oxygen because they have just used their oxygen in aerobic respiration to release energy from glucose. So the oxygen moves from a high concentration in the blood to a lower concentration in these cells.

Carbon dioxide diffuses in the reverse direction. It is produced during respiration in the tissues and organs and so is in a higher concentration within them. The blood moving towards the tissues and organs has a low concentration of carbon dioxide because it has just come from the lungs, where it unloaded carbon dioxide and picked up oxygen. So in body cells carbon dioxide diffuses from a high concentration to a low concentration in the blood. The blood now has a high concentration of carbon dioxide. It is transported to the lungs, where the carbon dioxide diffuses from an area of high concentration in the blood to an area of lower concentration in the alveoli. You then breathe it out.
Adaptations of the lungs

The combined surface area of your lungs is the total area that is open to air inside your lungs and to blood on the other side. Your lungs have a surface area of about half the size of a tennis court. This means they have a huge area to allow oxygen to diffuse from the alveoli into the blood and carbon dioxide to diffuse from the blood into the alveoli.

In addition to having a large surface area, your lungs are adapted for effective gas exchange by:

- having moist membranes that allow substances to diffuse faster across them
- alveoli and capillaries having thin linings (usually one cell layer thick)
- having a rich blood supply
- breathing (ventilation), providing the lungs with a regular supply of fresh air and removing air low in oxygen and high in carbon dioxide.

Diffusion also occurs in other places in the body.

Some of your cells make urea as a waste product. This is at high concentration in your cells and a lower concentration in your blood, so it diffuses from your cells to your blood. It is transported in the blood to your kidneys for excretion.

Diffusion in other organisms

The size of many organisms is determined by the maximum distance that substances can diffuse quickly. Insects, for example, do not have lungs and therefore do not breathe. They simply have a number of small tubes that run into their bodies. Oxygen diffuses from these tubes into the cells of the insect because the cells are using it for aerobic respiration. So it moves from an area of higher concentration in the tubes to an area of lower concentration in the cells. The maximum size of insects is in part determined by the distance that oxygen can quickly diffuse into their cells.
These smaller organisms do have an advantage over larger ones, however. **The smaller they are the greater the relative size of their surface area compared to their volume.** That is, they have a greater surface area to volume ratio. Large surface area to volume ratios in smaller organisms make it easier for them to get the oxygen they need (and get rid of carbon dioxide).

Large organisms like us need specialised exchange surfaces for exchange – such as alveoli in our lungs and villi in our intestines (see Chapter 4) – and a transport system such as the blood to transport substances around our bodies. This is because larger organisms have a smaller surface area to volume ratio.

Fish absorb dissolved oxygen into their blood by diffusion in their gills. These structures have a large surface area to maximise this.

> **Factors that affect diffusion**

**Concentration gradient**

The steeper the concentration gradient (the bigger the difference in the number of particles between an area of high concentration and an area of lower concentration), the more likely the particles are to diffuse down the concentration gradient. For example, the more deodorant you put on, the more the particles of deodorant are likely to diffuse into the air, and so the more likely other people are to smell them.

**Temperature**

At higher temperatures all particles have more kinetic energy. They move faster as a consequence. This means that they are more likely to spread out from their high concentration to areas of lower concentration.

**Surface area of the membrane**

The larger the surface area of the membrane the more particles can diffuse at once. Many people who have smoked for long periods of time have a reduced surface area in their lungs. This is called emphysema. Because their lung surface area is reduced, they are less able to get oxygen into their blood. They therefore often find it harder to exercise.
Activity

Investigating surface area to volume ratio and diffusion

A class carried out an investigation to examine the effect of surface area on the diffusion of dye. They were provided with three cubes of clear agar jelly that had been cut to different sizes (Figure 3.6).

Cube A was 1 × 1 × 1 cm.
Cube B was 2 × 2 × 2 cm.
Cube C was 4 × 4 × 4 cm.

Each cube was placed in a 200 cm³ beaker and the beaker was filled with 150 cm³ of blue dye. The cubes were left in the dye for 5 minutes. After this time the cubes were removed and any excess dye washed off before drying with a paper towel.

The cubes were then cut in half and observations were made on how far the dye had moved into the agar (Figure 3.7).

Questions

1. Give two variables that were controlled in this investigation.
2. To work out the total surface area (SA) for cube A, first the surface area of one face needs to be calculated: this is 1 × 1. This then needs to be multiplied by the total number of faces (6), so the calculation is 1 × 1 × 6 = 6 cm². To work out the volume, all the dimensions should be multiplied: 1 × 1 × 1 = 1 cm³. Copy and complete Table 3.1 by working out the surface area and volume for cubes B and C.

<table>
<thead>
<tr>
<th>Cube</th>
<th>Total surface area in cm²</th>
<th>Total volume in cm³</th>
<th>Surface area/volume</th>
<th>SA : V</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>6</td>
<td>1</td>
<td>6</td>
<td>6 : 1</td>
</tr>
<tr>
<td>B</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>C</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

3. Calculate the surface area divided by volume for cells B and C and add this to the table.
4. Use the values you worked out for Question 2 to complete the surface area to volume ratios (SA : V) for cubes B and C.
5. In which cube had the greatest proportion been dyed blue?
6. Explain how the dye entered each cube.

To make sure diffusion rates are fast enough, multicellular organisms have many adaptations to increase diffusion. Examples can be found in the lungs in humans, root hair cells and leaves in plants, and gills in fish. Gills contain many finely divided sections of tissue that are rich in blood capillaries. All the finely divided sections added together give a very large surface area.

Osmosis

We learnt in the previous section that particles of gases and liquids naturally move from areas of high concentration to areas of lower concentration by diffusion and that this can be across a membrane. Osmosis is the net diffusion of water from an area of high concentration of water (dilute solution) to an area of lower concentration of water (concentrated solution) across a partially permeable membrane. Water is the only substance that has a special...
name for diffusion. As in diffusion, no additional energy is used, and so this is a **passive process**. Because osmosis is from a high to a lower concentration of water, we say it is down a **concentration gradient**.

**Example of osmosis**

When it rains, water is present in a high concentration in the soil surrounding plant roots. The concentration of water inside the plant is lower, particularly if it hasn’t rained for a while. So the water moves naturally from the soil into the plant cells across the membranes of the cells by diffusion. Because this is water and it moves across a membrane to get into the cells, we call this process **osmosis**.

The water will then be carried in the xylem up to the leaves, where most of it will evaporate into the air through stomata (tiny pores). This process is called transpiration. Because water is continuously evaporating, it will continuously be ‘pulled up’ from the roots, which means that the root cells almost always have a lower concentration of water than in the soil, therefore allowing water to continue to enter the plant by osmosis.

### Comparing water concentrations

If two solutions have the same concentrations of water and solutes (substances in the water), then there is no net overall movement of water. The same volume of water will move in both directions if they are separated by a partially permeable membrane. We say these solutions are **isotonic**.

If one solution has a higher concentration of solute than another we describe the first one as being **hypertonic** to the other one. Crucially, because this has a higher solute concentration it has a lower water concentration. So if we took a red blood cell and put it into a very salty hypertonic solution (brine), the water from inside the blood cell would pass into the solution by osmosis. It would move from an area of high water concentration (inside the cell) to an area of lower water concentration (in the brine). The red blood cell would **shrink up** as a result.

The reverse is also true. If one solution has a lower concentration of solute than another, we describe it as being **hypotonic** to the other solution. Crucially, because this has a lower solute concentration it has a higher water concentration. So if we took a red blood cell and put it into pure water (a hypotonic solution) the water from outside the blood cell would move into it. It would move from an area of high water concentration (outside the cell) to an area of lower water concentration (inside the cell). The red blood cell would **swell up** as a result.
**Test yourself**

4 Define the term ‘osmosis’.
5 Name the plant cell that is adapted to allow plants to absorb water.
6 Describe one key difference between osmosis and diffusion.

**Investigate the effect of a range of concentrations of salt or sugar solutions on the mass of plant tissue**

Here is one way to investigate osmosis in potatoes, but your teacher may suggest you use another method or investigate different types of vegetables.

**Method**

1 Label six boiling tubes 0.0, 0.2, 0.4, 0.6, 0.8 and 1.0.
2 Using the volumes given in Table 3.2 and the 1.0 M solution of salt or sugar you have been provided with, make up the following concentrations in each boiling tube.

<table>
<thead>
<tr>
<th>Concentration in M</th>
<th>Volume of 1 M salt or sugar solution in cm³</th>
<th>Volume of distilled water in cm³</th>
<th>Total volume in cm³</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0</td>
<td>0</td>
<td>25</td>
<td>25</td>
</tr>
<tr>
<td>0.2</td>
<td>5</td>
<td>20</td>
<td>25</td>
</tr>
<tr>
<td>0.4</td>
<td>10</td>
<td>15</td>
<td>25</td>
</tr>
<tr>
<td>0.6</td>
<td>15</td>
<td>10</td>
<td>25</td>
</tr>
<tr>
<td>0.8</td>
<td>20</td>
<td>5</td>
<td>25</td>
</tr>
<tr>
<td>1.0</td>
<td>25</td>
<td>0</td>
<td>25</td>
</tr>
</tbody>
</table>

3 Using a chipper or corer, remove tissue from the middle of a potato and cut it into six equal 1 cm-long pieces.
4 Pat the first tissue sample dry with a paper towel, then measure and record its starting mass in a table like the one shown on the right.
5 Place a 1 cm-long piece of potato in the tube labelled 0.0 M and start the stopwatch.
6 Repeat this for the other five concentrations.
7 After 30 minutes (or the time specified by your teacher), remove the potato piece from the tube and dry it gently using a paper towel.
8 Record the end mass for the potato piece.

9 Repeat for the other concentrations after each sample has been in the solution for 30 minutes.

**Table 3.3** The results of an investigation into the effects of solute concentration on plant tissue.

<table>
<thead>
<tr>
<th>Concentration in M</th>
<th>Starting mass in g</th>
<th>End mass in g</th>
<th>Change in mass in g</th>
<th>Change in mass in %</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.2</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.4</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.6</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.8</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>1.0</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Questions**

1 Work out the change in mass for each potato piece and record it in your table. The mass changes will be positive if the potato piece got heavier and negative if it became lighter. Ensure you have clearly indicated this.
2 Calculate the percentage mass change for each piece of potato using the equation:

\[
\% \text{ change in mass} = \frac{\text{change in mass}}{\text{starting mass}} \times 100
\]

3 Now plot a graph of your data with the sugar or salt concentration on the x-axis and the percentage change in mass on the y-axis. Think carefully about how you will set up your axes to show both positive and negative values on the same graph.
4 Why did you have to pat dry the potato piece before and after each experiment?
Active transport

On occasions organisms need to move particles from areas where they are in low concentration to areas of higher concentration across membranes. This is called **active transport**, and we say that the particles are moving **up (or against) a concentration gradient**. If this is the case, then **energy** must be used. This is not a passive process like diffusion and osmosis. The energy needed comes from respiration.

**Examples of active transport**

**Mineral ions and plant roots**

We learnt in the previous section that water moves from an area of high concentration in soil to a lower concentration in plant roots across a membrane and that this is called osmosis. But plants need to take up **mineral ions** from the soil as well. These exist in very low concentrations in the soil but in high concentrations in the plants. So we might expect the mineral ions to diffuse out from the plant roots into the soil. Because the plants need to move the mineral ions from low to high concentrations, against the concentration gradient, they need to use energy. This is an example of active transport.

![Figure 3.11 Active transport of nitrates in an epidermal cell (root hair cell).](image)
Sugars and the digestive system

Following a sugary meal you will have high concentrations of sugars in your small intestine and lower concentrations in your blood. This means that sugars will naturally diffuse into your blood. But what happens if your last meal didn’t have much sugar in it? The lining of the small intestine is able to use energy to move sugars from lower concentrations in your gut into your blood.

![Figure 3.12](image)

▲ Figure 3.12 Look carefully at the nutrient concentrations in the intestine and the blood: (a) diffusion; (b) active transport.

### Test yourself

1. Define the term ‘active transport’.
2. Give one example of where active transport occurs in the human body.
3. Describe the key difference between diffusion and active transport in terms of the concentration gradient.

### Show you can...

Explain where active transport occurs in a plant.
Chapter review questions

1. Define the term ‘diffusion’.
2. Suggest an everyday example of diffusion of gases.
3. In which direction does oxygen diffuse in the lungs?
4. In which direction does carbon dioxide diffuse in the lungs?
5. Name the blood vessels from which oxygen diffuses into cells.
6. Give the scientific name for breathing.
7. Name the process by which oxygen moves into the blood from the lungs.
8. Define the term ‘osmosis’.
9. Describe where and how osmosis occurs in a plant.
10. Name the tiny holes in leaves.
11. Define the term ‘active transport’.

12. Is moving up a concentration gradient going from high to lower or from low to higher concentration?
14. Describe two ways in which your lungs are adapted for gas exchange.
15. Define the term ‘partially permeable membrane’.
16. Explain why mineral ions moving into a plant root is not an example of osmosis.
17. Describe what would happen to the size of a red blood cell if it were placed into a solution with the same concentration of solutes.
18. Describe what would happen to the size of a red blood cell if it were placed into a solution with a higher concentration of solutes.
19. Describe what would happen to the size of a red blood cell if it were placed into a solution with a lower concentration of solutes.
20. Describe one place where active transport occurs in plants.
21. Describe one place where active transport occurs in humans.

22. Explain why we say ‘net movement’ in our definition of diffusion.
23. Explain, in terms of diffusion, why insects are small.
24. Describe how you could use your hand and a length of string to model increasing surface area.
25. Explain how temperature affects diffusion.
26. Explain how the surface area of the membrane affects diffusion through it.
27. Describe an experiment in which you could investigate osmosis in plants using pieces of potato.
Practice questions

1 Figure 3.13 shows an alveolus and blood capillary in the lung.

![Figure 3.13](image)

a) During gas exchange, oxygen and carbon dioxide are exchanged between the alveolus and capillary. Which arrow (A or B) shows the net direction in which oxygen moves? [1 mark]

b) Gases move across cell membranes by diffusion.

i) Define the term ‘diffusion’. [2 marks]

ii) Copy and complete the sentence using words from the box below:

<table>
<thead>
<tr>
<th>active</th>
<th>passive</th>
<th>energetic</th>
<th>kinetic</th>
</tr>
</thead>
<tbody>
<tr>
<td>oxygen</td>
<td>energy</td>
<td>carbon dioxide</td>
<td></td>
</tr>
</tbody>
</table>

Diffusion is a ____________________________ processes. This means it does not require additional ____________________________. [2 marks]

2 Figure 3.14 shows three model cells (the pink areas) containing and surrounded by the same particles which can move freely into and out of the cell.

a) Which cell will have the greatest net movement of particles into it? [1 mark]

b) i) What effect would increasing the temperature have on the rate of movement of particles? [1 mark]

ii) Why would this occur? [1 mark]

![Figure 3.14](image)

3 Figure 3.15 shows a plant cell before and after it was placed in distilled water for 10 minutes.

![Figure 3.15](image)

a) Describe one way in which the cell looks different after being left in distilled water. [1 mark]

b) i) Describe what would happen if an animal cell were placed in the distilled water. [1 mark]

ii) Give the reason why this is different to what happened to the plant cell. [1 mark]

b) i) Draw a diagram to show what the plant cell would look like if it had been placed in a concentrated salt solution. [3 marks]

ii) Explain why the cell would look this way. [3 marks]

4 A student investigated the effect of osmosis on potato pieces. This is the method used:

a) The student started out with three pieces of potato that each measured 2 cm in length.

b) They then placed one piece of potato into a concentrated salt solution and one piece of potato into a dilute salt solution.

c) They left the potato pieces for 10 minutes.

d) They then removed the potato pieces and re-measured their length.

Describe how this method could be improved to produce valid results. [6 marks]
Presenting data in tables

Tables are an important part of most scientific investigations and are used to record the data collected. A good scientific table should present the data in a simple, neat way that is easy to understand.

When drawing tables there are some conventions (rules) to be followed:

► The independent variable (variable that is changed) should always be recorded in the first column. The dependent variable (variable that is measured) can be recorded in the next columns, with additional columns added if repeats are taken.
► The independent variable should be organised with an increasing trend.
► If a mean is calculated, this should be in the column furthest to the right.
► Column headers must have a clear title. If quantitative data are recorded, correct SI units must be given.
► Units must be given in the headers and not rewritten in the table body.
► All data in a column must be recorded to the same unit as the header, and mixed units should not be used.
► Data should be recorded to the same number of decimal places or significant figures.

Questions
1 Look at Table 3.4 and note down as many mistakes as you can see.

When this information to redraw Table 3.4 above so that it is correct.

3 Read the following instructions for an experiment examining osmosis in model cells. Draw a suitable table to record the volumes required in each beaker in order to prepare for the experiment.

Method for making up solutions

Collect five 100 cm$^3$ beakers and label them A, B, C, D and E. In each beaker add the following amounts of a concentrated fruit squash: A 100 cm$^3$, B 75 cm$^3$, C 50 cm$^3$, D 25 cm$^3$ and none in E. Then use distilled water to bring the volume of beakers B–E up to 100 cm$^3$. Stir the solutions to ensure they are mixed thoroughly.

Questions
4 Read through the method of the experiment on the next page and design a suitable table to record the results of the experiment. Ensure you identify the independent and dependent variables.
**Method**

Take five equal-sized pieces of Visking tubing that have been soaked in water. Tie one end of each securely. Using a pipette add 10 cm$^3$ of the solution from beaker A into one piece of Visking tubing. Tie the other end of the tubing using string and ensure that no liquid can escape. Repeat this process for the other four solutions B–E. Use a balance to determine the starting mass of each tube.

Place each tube in a separate beaker containing 200 cm$^3$ of distilled water. After 5 minutes remove the tubes and pat dry with a paper towel. Record the mass of each tube. Return to the beakers they came from and repeat, recording the mass at 10, 15 and 20 minutes.

**Figure 3.16** The liquid in the beaker is pure water. The red liquid is a very concentrated sugar solution with some red food dye added and is placed inside Visking tubing. This special tubing allows molecules of water through it but not larger sugar molecules. Water moves by osmosis through the Visking tubing from an area of high water concentration in the beaker to an area of lower water concentration (because of the added sugar) in the Visking tubing. This makes the red solution rise up the glass tube.
Animal tissues, organs and organ systems

Have you ever thought of yourself as a complicated tube with your mouth at one end and your anus at the other? This is one way you could look at your digestive system. Your digestive system breaks down food into pieces that can be absorbed into your body and used by it. Your circulatory system transports this food, and other crucial substances such as oxygen, in the blood. This is only possible because a muscular organ called the heart is able to pump the blood around the body, through more than 50,000 miles of blood vessels.

This chapter covers specification points 4.2.2.1 to 4.2.2.7 and is called Animal tissues, organs and organ systems.

It covers organisational hierarchy, the principles of organisation, the human digestive system and its enzymes, the heart and vessels, blood, related health issues, the effects of lifestyle, and cancer.
Levels of organisation in living organisms

In multicellular organisms, there are a number of levels of organisation. For example, in animals there are:

- **cells**: the basic building blocks of all living organisms
- **tissues**: groups of cells with similar structure and functions
- **organs**: groups of tissues that perform a specific function
- **organ systems**: organs are organised into organ systems
- **organisms**: the different organ systems make up organisms.

Table 4.1 gives some examples of these terms.

<table>
<thead>
<tr>
<th>Organisational level</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cell</td>
<td>Nerve cell, muscle cell</td>
</tr>
<tr>
<td>Tissue</td>
<td>Nervous tissue, skin</td>
</tr>
<tr>
<td>Organ</td>
<td>Brain, heart</td>
</tr>
<tr>
<td>Organ system</td>
<td>Nervous system, digestive system</td>
</tr>
<tr>
<td>Organism</td>
<td>Human, frog</td>
</tr>
</tbody>
</table>

The human digestive system

You eat large lumps of **insoluble** food. The breakdown products of this food must be transported around your body to reach the cells in which they are needed to complete life processes such as aerobic respiration. For this reaction alone each of your cells needs glucose (which comes from your food) and oxygen (which comes from your lungs). Digestion is the breakdown of food into smaller **soluble** pieces that can diffuse into your blood. The digestive system is the organ system that is responsible for doing this.
Functions of the parts of the digestive system

Salivary glands

Your salivary glands are found in your mouth and they make saliva, particularly when you are hungry and sense food. Saliva acts as a lubricant, making it easier to swallow food. It also contains an enzyme called amylase, which is an example of a carbohydrate enzyme as it breaks down starch (a carbohydrate) into simple sugar (glucose). When food mixes with saliva in your mouth the chemical breakdown (digestion) begins.

Oesophagus

Your mouth is connected to your stomach by a thin tube approximately 20 cm long. The only function of this tube, which is called the oesophagus, is to move food quickly and easily to your stomach. The saliva helps with this. The oesophagus is sometimes called the ‘food tube’ or ‘gullet’.

Stomach

The stomach is a small organ found between the oesophagus and the small intestine. It releases a type of enzyme called protease. This starts the chemical breakdown of protein. The stomach also releases acid. This has a pH of about 2 to 3. Stomach acid does not break down food. Instead it reduces the pH of the stomach to the optimum (best) level for protease enzymes to work properly. It also is part of our first line of defence against infection, as it destroys any pathogens that may have entered the body with food or water.
The human digestive system

KEY TERMS

**Bile** A green-coloured liquid produced by your liver, stored by your gall bladder and released into your small intestine to help break down fats.

**Lipids** Fats or oils, which are insoluble in water.

Liver

The liver is a large organ found to the right of your stomach. It produces a green liquid called **bile**, which helps to break down fats. Bile is not an enzyme. After being made in the liver, bile is stored in the **gall bladder** before being released into the small intestine.

Food does not actually pass into the liver. It moves from the stomach to the small intestine.

Pancreas

The pancreas is an organ that produces the three types of enzyme found in the digestive system: **carbohydrase** enzymes, which break down carbohydrates; **protease** enzymes, which break down proteins; and **lipase** enzymes, which break down **lipids** (fats). The pancreas releases these into the top section of the small intestine, in an area called the duodenum, close to where the gall bladder releases bile.

As with the liver and the gall bladder, food does not actually pass into the pancreas. It moves from the stomach to the small intestine.

Small intestine

Despite its name, the small intestine is actually the longest single part of the digestive system. It is called the small intestine because it is narrower than the large intestine. It is about 7 metres long and is responsible for **absorbing** the products of digestion into the blood. They are then transported around the body in the blood to where they are needed. It is adapted by being **folded** and having **villi**, both of which increase the surface area.

Villi

Villi (Figure 4.4) are microscopic finger-like projections of the lining of your small intestine. In an area about the size of your thumbnail you will have about 4000 villi. They massively increase the surface area of the small intestine and allow much more digested food to be absorbed into your blood. Each tiny villus contains blood capillaries, providing a rich blood supply to move digested food molecules to other parts of your body.

Peristalsis

The walls of the digestive system have rings of muscle around them and all along their length. These contract to squash lumps of food called boluses along your digestive system. **Peristalsis** is the rhythmic contraction of this muscle behind a bolus to push it along. This happens in the oesophagus and the small intestine.

Large intestine

Food that enters the large intestine is mainly indigestible fibre and water. The large intestine is wider than the small intestine but also much shorter, at a length of about 1.5 metres long. The large intestine is responsible for **absorbing water** (and salts) from the remaining digested food.
KEY TERMS

Substrate The molecule or molecules on which an enzyme acts.

Product The substance or substances produced by an enzyme reaction.

Table 4.2 The three types of enzyme found in the digestive system, the food groups they act upon and the molecules they are digested into.

<table>
<thead>
<tr>
<th>Enzyme</th>
<th>Substrate</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbohydrase</td>
<td>Carbohydrate</td>
<td>Sugars</td>
</tr>
<tr>
<td>Protease</td>
<td>Proteins</td>
<td>Amino acids</td>
</tr>
<tr>
<td>Lipase</td>
<td>Lipids (fats and oils)</td>
<td>Fatty acids (three molecules) and glycerol (one molecule)</td>
</tr>
</tbody>
</table>
Bile and lipase enzymes

As we have already seen, bile is produced by the liver and stored in the gall bladder before being released into the small intestine. It is not an enzyme but does help break down large globules of fat into tiny droplets. Any substance that does this is called an emulsifier, and the process is called emulsification. Bile does not actually break down the fat itself – instead it increases the surface area of the fat for the lipase enzymes to digest the fat into fatty acids and glycerol more quickly.

Bile is alkaline and so also neutralises hydrochloric acid entering the small intestine from the stomach. This increases the pH back towards neutral for the enzymes in the small intestine to work at their optimum.

Locations of enzymes in the digestive system

**Figure 4.7** The breakdown of complex food molecules into small, soluble, usable molecules.

**Figure 4.8** Digestive enzymes control reactions that take place in the digestive system. No digestive enzymes are made or used in the oesophagus, liver (bile is not an enzyme), gall bladder, large intestine or anus.
Use qualitative reagents to test for a range of carbohydrates, lipids and proteins

In this practical you will carry out tests to identify starch, sugars, lipids and proteins.

Your teacher will provide you with five samples labelled A–E. You need to determine which is a starch solution, which is a glucose solution, which is a protein solution, which is a lipid oil and which is water.

Copy the table below and use to collect your results:

Table 4.3

<table>
<thead>
<tr>
<th>Tube</th>
<th>Observation with starch test</th>
<th>Starch present?</th>
<th>Observation with Benedict’s test</th>
<th>Glucose present?</th>
<th>Observation with Biuret test</th>
<th>Protein present?</th>
<th>Observation with emulsification test</th>
<th>Lipids present?</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>B</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>C</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>D</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>E</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Use eye protection throughout the experiment.

**Testing for starch: method**
1. Using a pipette, add two drops of solution A into a well of a spotting tile.
2. Add two drops of iodine solution to this and record the colour observed.
3. Repeat this for the other four solutions.
4. If starch is present, a blue-black colour will be produced. Use your results to determine which tube contained starch.

**Testing for glucose: method**
1. Add 1 cm$^3$ of solution A to a boiling tube.
2. Add 10 drops or 1 cm$^3$ of Benedict’s reagent to this.
3. Place in a hot water bath (around 80 °C) and leave for 5 minutes.
4. Record the colour observed in your table.
5. Repeat this for the other four solutions.
6. If glucose is present, a brick-red precipitate will form. If it is not present, the solution will remain the blue of the Benedict’s reagent. Use your results to determine which tube contained glucose.

**Testing for protein (Biuret test): method**
1. Add 2 cm$^3$ of solution A to a test tube.
2. Add 2 cm$^3$ of Biuret solution to this.
3. Record the colour change in your table.
4. Repeat this for the other four solutions.
5. If protein is present, the solution will turn a light lilac purple colour. If it is not present, the solution will be a cloudy blue. Use your results to determine which tube contained protein.

**Testing for oil lipids: method**
1. Half fill a test tube with water.
2. Add one drop of solution A to this.
3. Move the test tube from side to side to mix thoroughly.
4. Place your thumb over the top of the test tube and shake.
5. Repeat this for the other four solutions.
6. As oils do not dissolve in water, an emulsion will form. This will make the water go cloudy if lipids are present.
The lock and key hypothesis

A hypothesis is a proposed explanation for scientific observations. The **lock and key hypothesis** explains how enzymes are specific for their substrates, just like a key is specific for the lock it fits. In the previous section you learnt that carbohydrase enzymes break down carbohydrate substrates and don't digest fats, for example. The lock and key hypothesis is a model that explains why.

Digestive enzymes are specific to the substrates they help break down because of their shape. Enzymes are proteins and all the proteins found in the body have a very specific shape to help them function.

A part of each enzyme is called its **active site**. This is the part that the substrate fits into. Any change in shape of an active site means that the breakdown will occur more slowly, or not at all. The enzyme and the substrate collide and become attached at the active site. The digestive enzyme then breaks the bonds holding the substrate together. Finally, the digestive enzyme releases the broken-down substrate.

Denaturing of enzymes

Enzymes work at an optimum **temperature** and **pH**. This is simply the temperature and pH at which they are most effective. Here we would say they have the highest enzyme activity. This is when the largest number of successful collisions takes place between the enzyme's active site and the substrate molecules. Any movement away from these optimum conditions will reduce the effectiveness of the reaction and so lower the enzyme activity.

Enzymes are **denatured** by high temperatures and extremes of pH. However, at low temperatures enzyme activity falls because of the lower kinetic energy, and if re-warmed the enzyme would start to work again. When denatured, enzymes are permanently damaged and won't work any more. The lock and key hypothesis can be used to explain this as well. When an enzyme is denatured, the shape of the active site has been altered so that it will no longer fit the substrate. In other words, the ‘key’ will no longer fit the ‘lock’.

**KEY TERMS**

- **Lock and key hypothesis** A model that explains the action of enzymes.
- **Active site** The region of an enzyme that binds to its substrate.

**Figure 4.9** How a digestive enzyme breaks down a substrate. Here, the substrate is sucrose and the products are glucose and fructose.

**KEY TERM**

- **Denatured** A permanent change to an enzyme as a result of extremes of pH or high temperature, which stop it working.

**TIP**

Avoid saying that enzymes have been ‘killed’. Use the term ‘denatured’ instead.
Investigate the effect of pH on the rate of reaction of amylase enzyme

In this practical you will examine how pH affects the digestion of starch by amylase.

Method
1. Add 5 ml of 1% starch solution and 5 ml of 1% amylase solution into separate boiling tubes.
2. Add 5 ml of pH buffer to the amylase in the boiling tube.
3. Place each boiling tube in a water bath at 30°C for 5 minutes.
4. After the 5 minutes add all the amylase and buffer into the boiling tube containing the starch solution and stir using a glass rod. Make sure you keep the starch/amylase/buffer mixture in the water bath throughout the investigation. The time when the amylase/buffer is added to the starch is time 0.
5. After 30 seconds remove a sample of the starch/amylase mixture and add two drops of this to iodine that had previously been added to a spotting tile.
6. Repeat every 30 seconds until the iodine solution remains yellow/brown.
7. Record the time taken.
8. Repeat at a range of pH values by using buffers of the chosen pH.

Questions
1. Why did you need to leave the tubes in the water bath for 5 minutes before mixing the amylase and the starch?
2. Plot a bar chart of your data with pH on the x-axis and time taken for the iodine to remain yellow-brown on the y-axis.
3. Use your bar chart to determine the optimum pH for amylase to work at.
4. Explain why the activity of amylase changes with pH.
5. Suggest how you could amend this investigation to improve the accuracy of the results.

Digestive (breakdown) and synthesis enzymes

So far you have learnt about the enzymes that are present in your digestive system. Digestive enzymes are those that break down (or digest) substances. Other enzymes, called synthesis enzymes, do the opposite (they help your body make complex molecules from simpler substances).

Examples of this are the enzymes found in protein synthesis. They join amino acids to make proteins. Enzymes also build up absorbed sugars into carbohydrates in the body and glycerol and fatty acids into fats.

Test yourself
4. Give an example of a carbohydrase enzyme.
5. Name the products of the breakdown of protein.
6. Describe how breakdown and synthesis enzymes are different.
7. Describe two ways in which enzymes can be denatured.

Required practical 4

Figure 4.10 Graph showing the effect of increase in temperature on an enzyme reaction.

Figure 4.11 The effect of increase in temperature (and denaturation) on an enzyme molecule. High temperatures alter the shape of the active site and so the ‘key’ doesn’t fit the ‘lock’.
The heart and blood vessels

Many substances need to be moved around your body. For example, for respiration, you need oxygen and glucose to be taken to all your cells. You need the waste products of this reaction, carbon dioxide and water, to be removed. Other substances such as hormones are also needed in specific organs at specific times. All of these substances travel in the blood pumped through blood vessels by the heart. Transport is the function of your circulatory system. It is composed of:

- the heart, which pumps the blood around the body
- blood, which carries the blood cells and key molecules around your body
- arteries, which carry blood from your heart
- veins, which carry blood back to your heart
- capillaries, which join arteries and veins through tissues and organs.

The heart is a pump which is responsible for pushing blood around your body. It is an organ made from muscle and nerve tissue. The muscle does the contracting and relaxing to push the blood around and the nerve tissue passes along electrical impulses to make sure the contractions happen correctly. The heart makes its own electrical impulses, which travel along its nervous tissue and cause the contractions. These electrical impulses are generated in the 'pacemaker' section, which is a small bunch of cells in the wall of the top right chamber (the right atrium). The pacemaker controls the rate of your heartbeat.
There are four chambers in your heart. The top two chambers are called the left and right atria (singular atrium). These collect the blood as it returns from your body. The bottom two chambers are called the left and right ventricles. The blood is pumped from the atria into the ventricles and then from the ventricles to the rest of the body. The blood on the left and right sides of the heart never mixes. The right ventricle pumps blood to the lungs (where it is oxygenated), and the left ventricle pumps oxygenated blood around the body. This is called a double circulation. In effect, the two sides of the heart pump blood to different places.

**Blood flow through the heart**

Blood returns from the lungs and is collected in the left atrium. Because it has come from the lungs it is high in oxygen and low in carbon dioxide. After entering the left ventricle, blood is pumped around the body. This pushes blood high in oxygen to all the tissues and organs that need it. The blood then is taken back to the right atrium by the vena cava. Because it has been to the tissues and organs, it now has low oxygen and high carbon dioxide levels. It enters the right ventricle and then is pumped to the lungs. Here diffusion removes the carbon dioxide and replenishes the oxygen. The blood then returns to where it began, the left atrium.
There are valves at the base of the arteries as they extend from the ventricles to prevent a backflow of blood. There are also valves between the atria and ventricles to stop blood being pumped backwards into the atria when the ventricles contract.

If you look closely at Figure 4.13 you will see that the walls of the left ventricle are thicker than in the right ventricle. This is because the left ventricle needs to pump the blood further to the extremities of your body, whereas the right ventricle only needs to pump it to the lungs, which are much closer.

**The blood vessels**

There are three types of blood vessel. Arteries move blood from the heart, veins take it back to the heart and capillaries carry it within tissues and organs. Capillaries link arteries and veins.

**Arteries**

Arteries must carry blood at high pressure, as it has just been pumped from the ventricles. Because of this they have thick walls made from elastic tissue and muscle tissue. This allows them to stretch. You can feel the surges of blood moving along your main arteries when you feel your pulse. You can do this in your wrist and neck, where arteries are particularly near the body surface.

The main artery coming from the left side of the heart, taking blood to the tissues and organs, is called the aorta. The main artery coming from the right side of the heart, taking blood to the lungs, is called the pulmonary artery. This is the only artery to carry deoxygenated blood.

**Veins**

Veins carry blood back to the heart at low pressure. This pressure has been lost as the blood travels through the arteries and capillaries. Veins also have to carry blood back to the heart against gravity from the lower parts of your body. Veins are wider than arteries but have much thinner walls. They have one-way valves to keep blood flowing in the correct direction. These are not present in arteries.

**TIPS**

- Arteries carry blood away from the heart.
- You only need to know the names of the aorta, pulmonary artery, pulmonary vein and vena cava.
Blood

Components of blood

Red blood cells

The red blood cells are what give our blood its red colour. In a cubic centimetre of blood there are approximately 5000 million red blood cells.
Red blood cells contain a substance called **haemoglobin**. This binds with the oxygen that diffuses into your blood in the alveoli. When it is carrying oxygen, it is called **oxyhaemoglobin**, and it turns the colour of the red blood cells from dark red to a brighter red. These red blood cells then move through arteries and capillaries to the organs and tissues that need the oxygen. Here oxygen diffuses from the red blood cell in a reverse of the reaction in the lungs.

Red blood cells are adapted for carrying oxygen in many ways. Their biconcave shape gives a high surface area to volume ratio, and having no nucleus means there is more room for haemoglobin.

**KEY TERMS**

**Haemoglobin** The protein in red blood cells that can temporarily bind with oxygen to carry it around your body.

**Oxyhaemoglobin** The name given to the substance formed when haemoglobin in your red blood cells temporarily binds with oxygen.

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**White blood cells**

White blood cells are part of your immune system, and they fight off invading pathogens (disease-causing **microorganisms**, such as bacteria). In a cubic centimetre of blood there are approximately 7.5 million white blood cells. Unlike red blood cells, white blood cells have a nucleus. There are two types of white blood cell. **Phagocytes** engulf pathogens and use enzymes to break them down. **Lymphocytes** produce **antibodies** to help clump pathogens together for phagocytes to destroy. **Antigens** are markers on the pathogens that the antibodies recognise and attach to.
Platelets

Platelets are **cell fragments**. In a cubic centimetre of blood there are approximately 350 million platelets. They are small structures that join together to form a scab when you cut yourself. Shortly after your skin is cut platelets **start the clotting process**. They do this by releasing chemicals called clotting factors. These turn a chemical called fibrinogen, which is found in your blood plasma, into fibrin. This forms a mesh and acts as a glue to help stick platelets together to form a scab.

Blood plasma

Plasma is a straw-coloured liquid that red and white blood cells and platelets are suspended in. It makes up about 55% of your blood and in turn is made from over 92% water. You have about 3 litres of blood plasma in the 5 litres of total blood. Many molecules that your cells need, such as glucose and amino acids, and those that are waste, such as carbon dioxide and urea, dissolve in your plasma.

**Test yourself**

12 Name the two types of white blood cell.
13 What is the colour of plasma?
14 Describe what happens when red blood cells meet oxygen in the lungs.
15 Describe how red blood cells are adapted to their function.

**Show you can...**

Describe the functions of the components of your blood.

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**Coronary heart disease: a non-communicable disease**

The heart is a large muscle that contracts to push blood through the blood vessels all around your body. But the muscle (and nerve) cells that make up the heart organ need to respire themselves to keep on contracting and relaxing. In order to do this they must be supplied with oxygenated blood. This comes through the **coronary arteries**. Glucose and oxygen diffuse from the blood in these arteries and their capillaries into the cells of the heart.

In **coronary heart disease**, layers of fatty material build up inside the coronary arteries, narrowing them. This reduces blood flow to the heart cells, resulting in a lack of oxygen for the heart muscle. High levels of **cholesterol** in the blood can lead to this happening. The fatty deposits slow or stop the oxygenated blood reaching the cells of the heart. Lack of oxygen can cause cells to die and eventually lead to a heart attack.
In recent years, coronary heart disease has become one of the major causes of death in the world. The traditional treatment for this was a heart bypass operation, in which a small section of artery is moved from another part of the patient’s body to short-circuit the blockage in the coronary artery. Now more patients are being treated by using less invasive stents. These are small devices made from mesh that are inserted into the arteries to keep them open. This operation is less dangerous and faster to recover from than a heart bypass.

Eating a balanced diet, stopping (or not starting) smoking, reducing alcohol intake, maintaining a healthy weight and regular exercise all reduce the risk of coronary heart disease. Drugs such as statins are also prescribed by doctors to reduce blood cholesterol, which in turn reduces the risk of coronary heart disease. Statins slow the rate at which fatty material is deposited in the arteries.

○ **Faulty valves**

The heart has four valves inside it to stop blood flowing backwards. Any flow of blood backwards reduces the efficiency at which blood flows around the body. This means that fewer glucose, oxygen and other molecules reach the cells that need them.

Valves that are faulty might not open properly or not close completely to stop backflow. This may cause breathlessness, tiredness, dizziness and chest pain for the patient. The most common form of treatment for severe cases is replacement of the valves during open-heart surgery. These valves can be replaced by valves from donors (biological valves) or artificial mechanical valves.

○ **Other heart problems**

Heart contractions are controlled by a bundle of cells called the pacemaker in the lining of the right atrium. These send electrical impulses down the heart’s nerve cells to regulate the contractions. Some people are born with or develop problems with these cells, which affect the timing of the electrical impulses. An artificial pacemaker can be fitted to take over the generation of electrical impulses.

○ **Transplants**

Heart failure is when a person’s heart stops beating. If this happens, or is likely to happen, a person usually requires a heart or heart-and-lungs transplant. These are the most serious of all operations described in this section. As with all transplants, a match between the donor and the patient must be found to stop the transplant being rejected by the patient’s immune system. Often patients are on long waiting lists until a suitable donor is found.

An artificial heart machine can sometimes be used in hospital to keep people alive when waiting for a heart transplant or when recovering from a heart operation. Essentially these are mechanical pumps that are connected to the patient’s blood supply.
The heart disease discussed in this section is an example of a non-communicable disease – a disease that is not infectious and cannot be passed from person to person.

Health issues

Health is defined as the state of physical and mental wellbeing. So being healthy means you are mentally as well as physically fit. Both physical and mental health can be maintained or improved by:

- a well-balanced diet
- regular exercise
- reducing stress
- seeking medical help for mental or physical difficulties.

**Well-balanced diet**

A well-balanced diet means that you have the correct amount of the key food groups. This is often shown in a food pyramid, as shown in Figure 4.28. Vegetables are low in fat, high in fibre and provide your body with key vitamins. Fruits have more natural sugar than vegetables do, but are also low in fat and high in fibre and vitamins. Fats should only be consumed in lower quantities and are found in fish and nuts as well as many processed foods. Dairy products include milk, yoghurt and cheese. These are high in protein and some vitamins but also high...
in fats and cholesterol. Recent research suggests that the negative effects of dairy foods can outweigh the benefits such as strengthening bones. Meat and beans are a good source of protein as well as vitamins and minerals. Some scientists think that the food pyramid is an oversimplification, whilst others think that it is a useful guide for the public.

### Regular exercise

The National Health Service (NHS) in the UK recommends that young people (aged 5 to 18) undertake at least 1 hour of physical activity every day. Some of this should be moderate intensity such as cycling and playground activities. Other activity should be vigorous, such as fast running and tennis. On 3 days a week this should involve muscle-strengthening exercise such as push-ups, and bone-strengthening activities such as running. Exercise also improves the effectiveness of your circulatory system.

### Physical and mental ill health

Diseases can cause ill health. Some different types of disease can interact to cause health problems. Problems with a person’s immune system might mean they are more likely to suffer from communicable diseases. A small number of specific virus infections can lead to the development of cancer. The reactions of a person’s immune system to infection from a pathogen can trigger allergies such as skin rashes and asthma. Severe physical ill health can lead to mental ill health, such as stress, anxiety and depression.

**Stress** is the feeling of being under too much mental or emotional pressure. This can affect how you feel, think and behave. It is common for people who are stressed to sleep badly, lose their appetite and have difficulties concentrating.

**Anxiety** is a feeling of unease, which might be worry or fear. This can be mild or severe depending upon the situation and the person. **Depression** affects different people in many different ways. Some people feel sad or hopeless, others lose interest in things they used to enjoy. Depression can also affect your physical health. It can make you feel tired and also lose your appetite. Severe depression can make people feel suicidal. People who feel stressed, anxious or depressed should speak to their doctor as soon as possible.

### Test yourself

18 Name the two types of wellbeing that make up health.
19 What should you do if you are worried about your health?
20 Describe the ways in which you can improve your physical or mental health.

### Show you can...

Describe the differences between stress, anxiety and depression.
Cancer

Sometimes cell differentiation or division goes wrong and cancerous cells are produced. These cells can divide rapidly by mitosis and quickly cause a lump or tumour. This rapid growth of cells is called uncontrolled cell division. There are two types of tumour. The first is described as malignant and causes cancer. These tumours divide rapidly and grow out of control. They can spread from one part of the body to another. If this happens it is called metastasis and forms secondary tumours. Prompt medical treatment is often needed to remove a malignant tumour or destroy the individual cells to stop the cancer spreading. The second type of tumour is described as benign and is medically less serious. Benign tumours are not cancerous, do not spread to other parts of the body and are usually contained within a membrane. They are often removed like malignant tumours, however. Nearly one in two people born after 1960 will suffer from cancer at some point in their life.

Signs of cancer include a lump formed by the tumour, unexplained bleeding, a long-term cough and a loss of weight without dieting. There are, however, over 200 different cancers in humans, and so there are many other symptoms not described here. The most important advice for anyone who thinks they might have cancer is to seek professional medical help as soon as they possibly can.

Screening

When doctors look for cancer it is called screening. This can be feeling a bump to see if it is a tumour, and taking blood tests, urine tests or X-ray images. Screening can also be undertaken before a person develops any symptoms, if they have a family history of developing cancer, for example.

Figure 4.30 An X-ray shows a tumour (central lighter zone) as being different to the surrounding tissue.
The effect of lifestyle on some non-communicable diseases

○ Causes of cancer
More than 20% of cancers are caused by smoking, and too much alcohol can also lead to cancer. Cancers can also be caused by infections such as hepatitis B and C, and the human papillomavirus (HPV). Cancer can also occur as a result of genetic disorders inherited from parents. Other factors such as ionising radiation (including the Sun’s ultraviolet rays) and environmental pollutants from industry are other causes. Obesity is a lifestyle factor that can contribute to cancer. The risk of cancer also increases as we age.

The most common cancers in the UK are breast cancer, lung cancer, prostate cancer (men only) and bowel (large intestine) cancer. It is very important to begin treatment of cancers as soon as possible. Sadly many cancers that are detected late become life-threatening. It is likely that some of these could have been treated if they were detected earlier.

○ Treating and preventing cancer
The two most common methods of treating cancer are chemotherapy and radiotherapy. Chemotherapy uses very powerful drugs to kill cancer cells. Radiotherapy uses X-rays to do the same thing. Both chemotherapy and radiotherapy cannot differentiate cancer cells from the other healthy cells around the tumour, so they can kill other nearby cells too.

Many cancers can be prevented from developing by leading a healthy lifestyle. This includes not smoking, not becoming over or underweight, not drinking too much alcohol, and eating healthily, including fresh fruit and vegetables.

Test yourself

21 What are the most common cancer types in the UK?
22 Name a cancer that can only occur in men and another that can only occur in women.
23 Describe the difference between malignant and benign tumours.
24 Describe how a number of cancers can be prevented.

KEY TERMS
Risk factor Any aspect of your lifestyle or substance in your body that increases the risk of a disease developing.
Causation The act of making something happen.
Correlation When an action and an outcome show a similar pattern but the action does not necessarily cause the outcome.

The risk of coronary heart disease increases with high blood pressure, smoking and excessive alcohol, high cholesterol and poor diet. Any aspects of your lifestyle or any substances found in your body or environment that are linked to the development of a disease are called risk factors. Some risk factors are proved to cause diseases (causation), while others are only linked to a higher chance of developing them (correlation). Examples of risk factors and their associated conditions (diseases) are shown in Table 4.4.
Many diseases are caused by the interaction of a number of factors. For example, people who smoke and drink excessive amounts of alcohol are more likely to be unfit and put on weight.

### Table 4.4 Risk factors, conditions (diseases) and their effects.

<table>
<thead>
<tr>
<th>Risk factor</th>
<th>Condition (disease)</th>
<th>Effects</th>
</tr>
</thead>
<tbody>
<tr>
<td>Diet, smoking and lack of exercise</td>
<td>Cardiovascular disease</td>
<td>Layers of fat build up inside coronary arteries, narrowing them.</td>
</tr>
<tr>
<td>Obesity and lack of exercise</td>
<td>Type 2 diabetes</td>
<td>Body cells do not respond to the hormone insulin, which helps control the glucose level in the blood.</td>
</tr>
<tr>
<td>Alcohol</td>
<td>Liver function</td>
<td>Long-term alcohol use causes liver cirrhosis. The cells in the liver stop working and are replaced by scar tissue. This stops the liver from removing toxins, storing glucose as glycogen and making bile.</td>
</tr>
<tr>
<td>Alcohol</td>
<td>Brain function</td>
<td>Excessive use of alcohol can also alter the chemicals in the brain (neurotransmitters), that pass messages between nerve cells. This can cause anxiety, depression and reduced brain function.</td>
</tr>
<tr>
<td>Smoking</td>
<td>Lung disease and cancer</td>
<td>Smoking can cause cancer in many parts of the body, including the lungs, mouth, nose, throat, liver and blood. It also increases the chances of having asthma, bronchitis and emphysema.</td>
</tr>
<tr>
<td>Smoking and alcohol</td>
<td>Underdevelopment of unborn babies</td>
<td>Alcohol and chemicals from cigarettes in the mother’s blood pass through the placenta to her baby. Without a fully developed liver the baby cannot detoxify these as well as the mother can. This can lead to miscarriage, premature birth, low birth weight and reduced brain function.</td>
</tr>
<tr>
<td>Carcinogens and ionising radiation</td>
<td>Cancer</td>
<td>Chemicals and radiation that cause cancer are called carcinogens. Tar in cigarettes, asbestos, ultraviolet from sunlight and X-rays are examples.</td>
</tr>
</tbody>
</table>

### Test yourself

25 Name the disease caused by carcinogens.
26 Name the two organs most likely to be damaged by long-term alcohol abuse.
27 Describe the effects of type 2 diabetes.

### Show you can...

Explain the difference between causation and correlation, giving an example in your answer.
Chapter review questions

1 Describe the function of saliva.
2 Where is bile stored?
3 Name the types of enzyme produced in the pancreas.
4 Give the function of the large intestine.
5 Name the organ that pumps your blood.
6 In what direction do arteries carry blood?
7 Explain why the term ‘double pump’ is used for the heart in mammals.
8 Explain why malignant tumours are more serious than benign ones.
9 Explain why doctors often use stents rather than transplants.

10 Explain why we must digest our food.
11 Define the term ‘enzyme’.
12 Name the enzyme that breaks down proteins, and the products.
13 Name the enzyme that breaks down carbohydrates, and the products.
14 Name the enzyme that breaks down fats, and the products.
15 Describe how you could use boiled and unboiled amylase to show that enzymes denature.
16 What are the two conditions that can denature an enzyme?
17 Define the term ‘optimum’ in relation to the effect of pH on enzyme activity.
18 Describe the pathway of the blood from the left atrium.
19 Describe how capillaries are adapted for their function.
20 Describe how phagocytes protect us from pathogens.
21 Name the blood vessels that provide the heart cells with glucose and oxygen.
22 Describe the effects of having faulty heart valves.
23 Describe one way in which doctors screen for cancer.
24 Define the term ‘anxiety’.

25 Explain how the lock and key hypothesis models enzyme action.
26 Name the part of an enzyme that is specific to the substrate.
27 Explain denaturing of enzymes using the lock and key hypothesis.
28 Suggest the effects of having a reduced platelet count.
29 Describe the causes of coronary artery disease.
30 Explain why doctors prefer to use stents than complete bypass operations.
Practice questions

1. Figure 4.32 below shows the main organs in the human digestive system.

   ![Figure 4.32](image)

   a) Name the following organs:
   i) organ A [1 mark]
   ii) organ B [1 mark]

   b) Organ C is the large intestine. What is its role? [1 mark]

   c) Figure 4.33 shows several villi. Villi are found within the digestive system.

   ![Figure 4.33](image)

   i) In which organ in Figure 4.32 would you expect to find the most villi? [1 mark]
   A Organ A    C Organ C
   B Organ B

   ii) From Figure 4.33, state one way the villus is adapted to maximise the absorption of the products of digestion. [2 marks]

   d) i) Give the term used to describe how food is moved through a digestive system. [1 mark]
   ii) Explain how this movement is brought about. [2 marks]

2. Which of the following would be the least invasive method of treatment for coronary heart disease? [1 mark]
   A Stent    C Pacemaker
   B Bypass   D Transplant

3. Figure 4.34 shows amylase speeding up the breakdown (digestion) of a large molecule.

   ![Figure 4.34](image)

   a) Why do large molecules need to be digested? [1 mark]
   b) What is region X in Figure 4.34? [1 mark]

4. Figure 4.35 below shows three types of blood vessel.

   ![Figure 4.35](image)

   a) Name the three blood vessels. [3 marks]

   b) Explain how the build-up of fatty material in the blood vessels that supply the heart can lead to a heart attack. [2 marks]

   c) Copy and complete Table 4.5 to show which part of blood fits the different descriptions listed. Choose your answers from the box. [4 marks]

<table>
<thead>
<tr>
<th>Description</th>
<th>Part of blood</th>
</tr>
</thead>
<tbody>
<tr>
<td>Contain a substance called haemoglobin, which</td>
<td>Red blood cells</td>
</tr>
<tr>
<td>binds with oxygen</td>
<td></td>
</tr>
<tr>
<td>Fight off invading pathogens</td>
<td></td>
</tr>
<tr>
<td>Small structures that can join together to</td>
<td></td>
</tr>
<tr>
<td>prevent blood loss</td>
<td></td>
</tr>
<tr>
<td>Carries many molecules such as glucose and</td>
<td></td>
</tr>
<tr>
<td>amino acids and carbon dioxide</td>
<td></td>
</tr>
<tr>
<td>Have a shape designed to maximise their</td>
<td></td>
</tr>
<tr>
<td>surface area</td>
<td></td>
</tr>
</tbody>
</table>

5. A student carried out food tests to look for the presence of glucose, starch and protein. Describe how they would have carried out the food tests.

   You should include:
   a) what reagents you would use [6 marks]
   b) what positive results would look like.
Understanding and evaluating models

The word ‘model’ is used a lot in science but what does it actually mean? Take a minute to try and define your understanding of the term ‘model’ and note down as many models you can think of that you have encountered in your science lessons.

Scientific models can take many forms, but their main purpose is to represent a process or feature in a way that is easier to predict, understand, visualise or test. Some models are scaled-up versions of real things, such as the model of DNA or a model cell. Others are scaled down, like the model of the solar system. Other models are much more abstract, explaining phenomena we can’t see and simplifying the details of them, such as the particle model. Although models are designed to help us understand by simplifying versions of reality, they can also be misleading.

As you have seen, the lock and key hypothesis is a model used to explain how enzymes work. This is shown in Figure 4.36.

Questions

1. What are the strengths of this model in representing how enzymes work?
2. What are the weaknesses of this model?
3. What does the model fail to explain or represent about enzymes?
Visking tubing is a membrane that has small holes in it. These holes are large enough to allow small molecules such as water and glucose to pass through but are too small to allow larger molecules such as starch to pass through. We can use this model to demonstrate how large molecules are digested into smaller molecules, which can leave the digestive system and enter the blood.

**Figure 4.37** An experiment to model the digestive system.

**Questions**

4. What does the Visking tubing, and the water surrounding it, represent in this model?

5. Explain why at the start of the experiment no glucose is present inside the Visking tubing or in the water in either beaker, but over time glucose is detected in the tubing and the water in beaker 1.

6. Explain why starch was only found inside the Visking tubing and not in the water.

7. What features of the digestive system does this model not represent appropriately?

Using Plasticine or other modelling materials, create models to show the key differences between the different types of blood vessel. Ensure you can explain how each blood vessel is adapted to carry out its function.

**Questions**

8. Evaluate your models and identify three strengths and three weaknesses.
Plants are truly amazing. Without them and other photosynthesising species it is possible that the only life on Earth would be a few organisms surrounding volcanic vents on the bottom of the ocean. Plants support almost all life on Earth including you and me. Plants have far more in common with us than you might think. Like us, they are complex organisms that are arranged into tissues, organs and systems. And they are highly adapted to live and reproduce in their natural environment, just like us.
Plant tissues

**Epidermal tissue**
The epidermis is a tissue made up of single layer of cells that forms the outer layer of a plant. It has many functions, including protecting against water loss, regulating the gases that are exchanged between the plant and the air (especially in the leaves), and water and mineral uptake (especially in the roots). The epidermis is usually transparent, possessing fewer green chloroplasts in its cells than other plant tissues. Because the epidermis is transparent, light can pass through it and reach the palisade mesophyll in leaves.

**Palisade mesophyll**
Immediately below the upper epidermis of plant leaves is the palisade mesophyll tissue. The cells in this tissue are often more tightly packed and have a more regular shape than the cells in the spongy mesophyll tissue below them. Palisade mesophyll cells have more chloroplasts than all other plant cells and so are the major site of photosynthesis. The palisade cells are highly adapted for photosynthesis.

**Spongy mesophyll**
Spongy mesophyll tissue is found below the palisade mesophyll tissue towards the lower surface of plant leaves. The cells in this tissue are more spherical in shape than palisade mesophyll cells and have many spaces between them. Gases enter leaves through tiny pores called stomata (singular, stoma; see Figure 5.1), which have guard cells around them. These cells control the size of the opening. More stomata are normally found on the underside of leaves than on the upper surface. Spongy mesophyll cells have a large surface area in contact with the air spaces in the leaf to maximise gas exchange.

**KEY TERMS**
- **Epidermis** The outermost layer of cells of a plant.
- **Palisade mesophyll** Tissue found towards the upper surface of leaves with lots of chloroplasts for photosynthesis.
- **Spongy mesophyll** Tissue found below the palisade layer(s) of leaves with spaces between them to allow gases to diffuse.

**TIP**
Guard cells of stomata are the only epidermal cells to have chloroplasts.

**Previously you could have learnt:**
- All multicellular organisms including plants are organised in particular ways with a hierarchical arrangement in the sequence: cells, tissues, organs, organ systems and organisms.
- Plants make carbohydrates in their leaves by photosynthesis and gain mineral nutrients and water from the soil via their roots.
- Leaf stomata are important structures for gas exchange in plants.
- As in animals, plants also reproduce to produce young.
- Leaves are adapted for photosynthesis.

**Test yourself on prior knowledge**
1. Where do plants absorb water and carbon dioxide?
2. Describe the role of stomata in gas exchange.
3. Explain how leaves are adapted for photosynthesis.
Observing stomata on leaves

Stomata are the pores in a leaf that allow gas exchange with the atmosphere, which is required for photosynthesis and respiration. The numbers and arrangement of stomata vary on the upper and lower surfaces of plants and between plant species.

Method
1. Select the plant you want to study and use a small paintbrush to coat a small section of the upper and lower epidermis of a leaf with a water-based varnish or nail polish.
2. Leave the varnish to dry fully.
3. Using tweezers and sticky tape remove the varnish impressions from the leaves and mount each onto a labelled slide.
4. Observe the stomata using a microscope.
5. Draw and label a sketch of a stoma and the surrounding guard cells.

Questions
1. On which side of the leaf was there a higher density of stomata? Suggest why this is the case.
2. Explain how you would expect the stomata to look in a plant that is wilted compared with one that has been recently watered.

Xylem and phloem

Water flows up through xylem tissue from the roots to the leaves during transpiration. Phloem cells carry the glucose (in the form of sucrose) made in photosynthesis from the leaves of a plant to all other parts of the plant during translocation. Xylem and phloem tissues are often found together in vascular bundles.

Meristem

The meristem is the region of plant tissue in which stem cells are produced and so where much of the plant’s growth occurs. They are found in shoot tips reaching for the sunlight and root tips following gravity downwards.
Plant organs

Root

Roots are plant organs that are usually found below the soil. As a result they are white because they don’t contain green chloroplasts for photosynthesis. Roots absorb water by osmosis and minerals by active transport from the soil. They also anchor the plant into the soil. In addition, in some plants, roots can store the glucose made during photosynthesis, usually as starch.

The meristem is found at the very tip of the root. Here new cells are produced to allow the root to grow deeper into the soil. On the outside of roots are root hair cells to absorb water by osmosis. These are specialised epidermal cells. In the middle of the root are the xylem and phloem tissues.

▲ Figure 5.5 Look how many root hairs are on this one tiny root.

Test yourself

1. Name the two tissues involved in plant transport.
2. Describe how palisade mesophyll cells are adapted to their function.
3. Describe how the arrangement of spongy mesophyll cells helps the plant.

Show you can...

Label the parts of the cross-section of a leaf in Figure 5.2.

TIP
Remember that organs are parts of an organism that have a particular function and are formed of different types of tissue.
Plant organ systems

Shoot

Scientists define a shoot as the stem, its leaves, and its buds (not just the very tip of a young plant).

The meristem is found at the very tip of the shoot. Here new cells are produced to allow the shoot to grow towards the light. On the outside of shoots are epidermal cells.

Leaf

The leaf is a plant organ and is the major site of photosynthesis. It also controls the flow of water through the plant. Previously you learnt that water is absorbed by osmosis from the soil into the roots. It is then ‘pulled’ through the plant by the transpiration stream because it is continuously being released from the leaves through stomata, which open and close to regulate this process.

Test yourself

4 Name the process by which water is absorbed into the roots.
5 What are the two functions of roots?
6 Where are meristems of shoots found?
7 Explain why roots are often white.

Transportation organ system

You have already learnt that xylem and phloem are tissues, and that roots, shoots and leaves are plant organs. These combine to make the plant transportation organ system, which transports all substances around a plant.

Figure 5.6 (a) A cross-section of a leaf showing the large surface area for gas exchange provided by the internal air spaces, and the movement of materials through the stomata. (b) Root hair cells give the roots a large surface area for absorption.
Transpiration and the transpiration stream

Water enters root hair cells in plant roots by osmosis. It then travels by osmosis through the cells of the root and then enters xylem cells. It travels up through the root and stem in long continuous columns of xylem cells. Eventually the xylem branches to form veins that carry the water to the leaves, where it enters the leaf cells.

Much of this water evaporates out of the leaf cells (mainly the spongy mesophyll cells) and enters the leaf air spaces as water vapour. This then diffuses out of the leaf through the air spaces and stomata. This is a continuous process, and the loss of water from a plant through the leaves is called transpiration. The constant evaporation of water from the leaves pulls, or ‘sucks’, the water up through the rest of the plant in a long, unbroken transpiration stream.

Transpiration has a number of functions, including:

- providing water for leaf cells and other cells (e.g. to keep them turgid)
- providing water to cells for the process of photosynthesis
- transporting minerals to the leaves.

Diffusion of any substance happens faster if the concentration gradient is greater (that is, the difference between the high and low concentrations is bigger). If the air surrounding a leaf is very humid (like just before a thunderstorm) then the water vapour gradient will be less steep so the rate of transpiration will be lower. On windy days the air surrounding the leaves is continually replaced. This keeps the concentration gradient steep and the rate of transpiration high. When temperatures are higher the rate of evaporation of water is higher and so transpiration occurs more rapidly. Water is also used up more rapidly during the daylight hours as some of it is used to make glucose by photosynthesis, so transpiration is increased. Also, the stomata are more likely to be open during the day.

In summary, high rates of transpiration are achieved when:

- there is more wind
- there is a high temperature
- the air is less humid
- the light intensity is high (during the day).

Translocation

Phloem tissue is also part of the transport organ system. Phloem transports dissolved sugars that are made in the leaves by photosynthesis to the rest of the plant. The transported sugar is usually either immediately used in respiration or stored as starch. The movement of dissolved food through the phloem is called translocation.
Investigating transpiration

A class were investigating water loss from plants and wanted to compare the amount of water lost from the upper and lower surfaces of a leaf. Four leaves of similar sizes were selected from a bush and their surface areas estimated by drawing around them on squared paper.

A thin layer of petroleum jelly was used to cover the stalks of the leaves and applied to the epidermises of some of the leaves, as shown in Table 5.1. The leaves were weighed and their starting mass recorded. They were then each hung on a piece of string using a paper clip attached to the stalk and left undisturbed in the light on a windowsill in the classroom. After an hour the leaves were re-weighed.

Table 5.1

<table>
<thead>
<tr>
<th>Treatment</th>
<th>Estimated surface area in cm²</th>
<th>Starting mass in g</th>
<th>End mass in g</th>
<th>Change in mass in g</th>
</tr>
</thead>
<tbody>
<tr>
<td>A No petroleum jelly</td>
<td>50</td>
<td>1.34</td>
<td>1.14</td>
<td></td>
</tr>
<tr>
<td>B Petroleum jelly on the upper surface</td>
<td>55</td>
<td>1.46</td>
<td>1.38</td>
<td></td>
</tr>
<tr>
<td>C Petroleum jelly on the lower surface</td>
<td>52</td>
<td>1.42</td>
<td>1.38</td>
<td></td>
</tr>
<tr>
<td>D Petroleum jelly on both sides</td>
<td>1.56</td>
<td>1.55</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Questions

1 Copy and complete Table 5.1 by estimating the surface area of leaf D in Figure 5.7.
2 Determine the change in mass for each leaf.
3 Give an explanation for the class’s results.
4 Calculate the water loss for leaf A in g per cm².
5 What was the independent variable in this experiment?
6 What was the dependent variable?
7 Name a variable that should have been controlled in this experiment but was not.

![Figure 5.7 Measuring the area of leaf D using squared paper.](image)

Test yourself

8 Name the plant organ from which water evaporates.
9 Name a plant organ system.
10 Describe the diffusion of oxygen from plant leaves.
### Chapter review questions

1. Describe where in a leaf the spongy mesophyll layer is found.
2. Describe an adaptation of the palisade mesophyll cells.
3. Name the tissue in which water moves up a plant from its roots.
4. Explain why most roots are white.
5. Describe two functions of roots.
6. Name the tissue in a plant that produces stem cells.
7. Name the process by which plants absorb mineral ions from the soil.
8. Give an example of a plant organ.
9. Name the type of cell in which most chloroplasts are found.
10. Name the organs that make up the plant transportation organ system.
11. Name the process by which water enters a plant root.
12. Describe the difference in structure of palisade mesophyll and spongy mesophyll tissues.
13. Describe the location and function of guard cells.
14. Describe an experiment in which you could use nail varnish to investigate the number of stomata on different plant leaves.
15. Name two tissues involved in transportation in a plant.
16. Suggest where you might find a plant with green roots.
17. Describe the process of transpiration.
18. Explain why more stomata are found on the lower surface of leaves.
19. Explain when you might expect all the stomata of a plant to be open.
20. Explain why plants must continuously allow water to evaporate from their leaves.
21. Explain how increasing humidity affects the rate of transpiration.
22. Explain how decreasing temperature affects the rate of transpiration.
23. Suggest when during the day transpiration is most likely to be highest.
24. Explain why more transpiration happens on a windy day.
25. Describe an experiment in which you could use petroleum jelly to investigate transpiration in leaves.
Practice questions

1 Figure 5.8 shows a cross-section part of a leaf.

\[ \text{Figure 5.8} \]

a) Copy and complete Table 5.2 by identifying tissues A–C. [3 marks]

<table>
<thead>
<tr>
<th>Table 5.2</th>
</tr>
</thead>
<tbody>
<tr>
<td>Tissue</td>
</tr>
<tr>
<td>Spongy mesophyll</td>
</tr>
<tr>
<td>Epidermis</td>
</tr>
<tr>
<td>Palisade mesophyll</td>
</tr>
</tbody>
</table>

b) Two other tissues found in plants are xylem and phloem. These are often found together in bundles and have an important role in transporting substances around plants. Name something that is transported by:

i) the xylem [1 mark]

ii) the phloem. [1 mark]

2 Figure 5.9 shows the arrangement of stomata on the underside of the leaves of two species of plant. Each square represents 0.02 mm$^2$ of leaf.

\[ \text{Figure 5.9} \]

a) Name cell X. [1 mark]

b) Do you think species A or B is adapted to live in a drier habitat? Explain your reason. [2 marks]

c) Suggest another adaptation that the leaves might have to help them survive in a dry habitat. [1 mark]

d) Calculate the number of stomata per 1 mm$^2$ of leaf epidermis for species B. Show your working. [2 marks]

3 Figure 5.10 shows an experiment involving water loss in plants.

\[ \text{Figure 5.10} \]

a) Flasks A and B were weighed and shown to have the same mass at the start of the experiment. After 10 minutes they were re-weighed. Suggest what you would expect the mass of flask A to be compared to that of flask B at the end of the experiment. [1 mark]

b) Why was flask B needed? [1 mark]

c) Suggest why cotton wool was placed in each flask. [1 mark]

d) What is the name of the process by which leaves lose water vapour? [1 mark]

A) Transportation

B) Transformation

C) Translocation

D) Transpiration

e) Describe how this process occurs in a plant. [3 marks]

f) In which of these conditions would water loss from a plant be greatest? [1 mark]

A) Hot and humid conditions

B) Cold and humid conditions

C) Hot and dry conditions

D) Cold and dry conditions
Understanding error

There are often differences in the results obtained in an experiment caused by different types of error. A random error is usually caused by a mistake being made by the person carrying out the experiment, a change in the measuring instrument or a change in the environment that was not controlled. Random errors cause the result to vary in an unpredictable way, spreading around the true value. We can reduce the effect of random error by carrying out more repeats and calculating a mean.

A systematic error causes the readings to differ from the true value by a consistent amount each time a measurement is made. These types of error usually come from the measuring instrument, either because it is incorrectly calibrated or because it is being used incorrectly by the experimenter. Systematic errors cannot be dealt with by more readings; instead the whole data collection needs to be repeated using a different technique or a different set of equipment.

Four students were asked to examine error by using a piece of equipment called a potometer. A potometer measures the rate of transpiration in plants. As water is lost from the leaves the plant draws up water to replace it. By measuring the distance moved by an air bubble over a set period of time the rate of transpiration can be measured. This allows transpiration rates under different conditions to be compared.

The students were first asked to set up their potometers and make a mark where the air bubble was at the start of the experiment. After 10 minutes they recorded the distance their bubble had moved. They repeated this process three times to determine a mean result.

Their results are shown in Table 5.3.

Table 5.3 The distance moved by an air bubble in a class potometer experiment.

<table>
<thead>
<tr>
<th>Student</th>
<th>Distance moved by the air bubble in cm</th>
<th>Repeat 1</th>
<th>Repeat 2</th>
<th>Repeat 3</th>
<th>Mean</th>
</tr>
</thead>
<tbody>
<tr>
<td>Amy</td>
<td></td>
<td>2.1</td>
<td>2.0</td>
<td>2.2</td>
<td>2.1</td>
</tr>
<tr>
<td>Chris</td>
<td></td>
<td>1.4</td>
<td>1.4</td>
<td>1.6</td>
<td>1.5</td>
</tr>
<tr>
<td>PJ</td>
<td></td>
<td>2.2</td>
<td>2.3</td>
<td>2.1</td>
<td>2.2</td>
</tr>
<tr>
<td>Kelly</td>
<td></td>
<td>1.4</td>
<td>2.8</td>
<td>2.4</td>
<td>2.2</td>
</tr>
</tbody>
</table>
Questions

1. Which set of data contains more errors? How do you know this from the data?

2. Chris's data were lower than those of the other members of the group, so the teacher asked him to show how he made his measurements. Figure 5.11 is a picture of how Chris recorded the distance moved, and how Amy did. What did Chris do wrong? What type of error is this? What should he now do?

3. The teacher asked Kelly to explain how she took her readings. She explained that she took the first reading sitting down, the second standing up, and the third level with the bubble. Why did her method for measuring the bubble introduce errors? What should she have done? Draw a diagram to explain why changing her position would have led to errors in her readings.

4. Why did the rubber tube have grease smeared around it? If this was not present how could it lead to errors in the data?

5. Identify any other potential sources of error in the experiment.

6. The students repeated their experiment with three different conditions:
   a) on a bench with a lamp shining on the plant
   b) on a bench with a fan blowing onto the plant
   c) on a bench with a plastic bag around the plant, creating a humid environment.

   For each condition, explain how you think the result would be different and why.

Figure 5.11 The results of an experiment using a potometer.
Infection and response

We catch communicable diseases from infected people. They are contagious. HIV/AIDS currently kills over a million people per year and estimates suggest that it has killed about 39 million people in total so far. History reveals much higher numbers of deaths from other communicable diseases. Smallpox is likely to have killed more than 300 million people in the 20th century alone. Spanish flu killed between 50 and 100 million in the same timescale. The Black Death killed around 75 million, reducing the world’s population to around 350 million in the 14th century.

This chapter covers specification points 4.3.1.1 to 4.3.1.9 and is called Infection and response.

It covers communicable diseases caused by viruses, bacteria, fungi and protists. It also covers human defence systems, vaccination, and the discovery and development of drugs, including antibiotics and painkillers.
Communicable (infectious) diseases

A pathogen is any microorganism that passes a communicable disease from one organism to another. There are four main types of microorganism that cause disease:

1. **viruses**, e.g. measles
2. **bacteria**, e.g. salmonella
3. **fungi**, e.g. rose black spot in plants
4. **protists**, e.g. malaria.

All four types of pathogen have a simple life cycle. They infect a host, reproduce (or replicate in the case of viruses), spread from their host and infect other organisms. This process then repeats. Many pathogens can survive without their host for a short period of time, but these are unable to reproduce without their host.

Pathogens are highly adapted to their role. They are very easily passed from one organism to another. We call these highly infectious. An example of a highly infectious pathogen is the measles virus. This is particularly easily spread because it is transmitted in the air. Other pathogens, such as the norovirus (winter vomiting bug), can reproduce very quickly. Others can survive for long periods without a host. An example of this is the *Staphylococcus* bacterium.

Test yourself

1. Give an example of a viral pathogen.
2. Give an example of a fungal pathogen.
3. Define the expression ‘highly infectious’.
4. Explain the advantage to a pathogen of reproducing quickly.

Show you can...

Describe the life cycle of an infecting pathogen.
Spread of pathogens

Pathogens have evolved many different ways of passing from one organism to another.

- **Airborne**: the common cold virus is often spread in tiny droplets of water propelled through the air when an infected person sneezes.
- **Through dirty water**: the cholera bacterium is often spread in unsterilised water.
- **By direct physical contact**: this can be sexual or non-sexual. Chlamydia is a bacterial pathogen that is one of the most common sexually transmitted diseases (STDs) in the world. Without treatment with antibiotics this can lead to serious reproductive problems.
- **Through contaminated food**: the *Escherichia coli* bacterium is often spread in uncooked or reheated food. It causes food poisoning.
- **Passed by another animal**: some farmers in the UK believe that badgers can catch the tuberculosis bacterium and pass it to their cattle. We call any organism that does this a **vector**.

**Test yourself**

5 Name the method by which cholera is often spread.
6 Describe how tuberculosis is spread.

**Show you can...**

Describe how the four different types of pathogen could be spread, giving an example of each in your answer.

Viral diseases

Viral diseases are those caused by a virus. They are the smallest pathogens and the most simple, made from a strand of genetic material (DNA or RNA) surrounded by a protein coat. They infect a cell in a host and use it to copy their genetic material and protein coats. These are then assembled into new virus particles. At this point cells are often full to bursting with new viruses, and they then split open to release new infecting particles to repeat the cycle.

- **Measles**

Measles is a highly infectious, common viral disease usually transmitted between young children. Its transmission is airborne, so it is passed in tiny droplets when an infected person sneezes. These are breathed in by those around them, who then may become
infection. Its symptoms include a fever and a red skin rash. In
developed countries babies are immunised (vaccinated) against
this infection. It can infect adults that were not immunised or did
not catch it as a child. More serious medical complications can
then occur including sterility in adults and foetal abnormalities in
pregnant women.

HIV/AIDS
HIV stands for human immunodeficiency virus. This is transmitted
when body fluids are shared, often during sex, or by shared use of
needles by drug users. It can be passed from mother to child in the
uterus, during birth or in breast milk. Immediately after infection,
the symptoms are like those of flu. After this, infected people usually
show no symptoms. In fact many may not know they are infected at all.
Months or years after infection, the virus attacks the body's immune
cells. The virus enters the lymph nodes and destroys a particular
type of immune cell. This means infected people are then less able
to fight off infections such as the tuberculosis bacterium, and they
are also more prone to a range of cancers. At this point the disease is
called late-stage HIV or acquired immune deficiency syndrome (AIDS).
There is currently no cure for HIV/AIDS. Infected people are given
antiretroviral drugs, which can slow the development of the disease.

Tobacco mosaic virus
The tobacco mosaic virus is a virus that affects the tobacco plant and
many other related species including tomatoes.

The leaves of infected plants develop a mottled or mosaic appearance
with the patches of normal tissue alternating with patches that have
a light green or yellow appearance, as shown in Figure 6.4. Leaves are
often wrinkled and curly.

The virus enters and damages chloroplasts in parts of the leaf, and this
explains the loss of green colour and the yellowing of leaves.

As the chloroplasts and leaves of infected plants are damaged, the
plants cannot photosynthesise as effectively and poor growth results.
This obviously affects the profit of the crop growers.

Tobacco mosaic virus infects plants by entering through parts that
have been damaged. The virus is easily spread from plant to plant on
tools and on the clothes or hands of farm workers. It can also spread
through contact between an uninfected and an infected plant (e.g. an
infected plant being blown against a neighbouring plant in the wind).
Occasionally, insects feeding on the plants may cause infection.

There is no cure, therefore the emphasis is on prevention. Measures to
reduce infection include:

- using strains that are resistant to the virus
- removing infected plants immediately the infection is spotted
- removing weeds in close proximity to the crops as they can harbour
  the virus.
Bacterial diseases

Bacterial diseases are caused by pathogenic bacteria. It is worth remembering that there are many useful bacteria that are not pathogens, including those that live in your digestive system and help you digest your food. All bacteria are prokaryotes. They are larger than viruses but are still only visible using a microscope. They live inside their hosts, often in mouths, noses and throats, but not inside cells as viruses do. Many pathogens produce toxins (poisons) as they grow, which irritate the surrounding cells of the host.

- **Salmonella**
  Salmonella is normally spread in food that has been prepared in unhygienic conditions, that has not been cooked well enough or that has been kept too long. We become infected if we eat food containing salmonella bacteria. In the UK, poultry are vaccinated against salmonella, to control its spread. Salmonella and its toxins cause fever, abdominal cramps, vomiting and diarrhoea in humans. Prevention measures include vaccinating poultry against salmonella, cooking food thoroughly and preparing food in hygienic conditions.

- **Gonorrhoea**
  Gonorrhoea is a sexually transmitted disease (STD) caused by a species of bacterium. Symptoms include a painful burning sensation when urinating and the production of a thick yellow or green fluid (discharge) from the vagina or penis. It is spread by sexual contact. Prevention measures include the use of a barrier method of contraception such as a condom, because this prevents the bacteria passing from person to person. Gonorrhoea can be treated with antibiotics, although some resistant strains of bacteria have evolved, making treatment more difficult.

**Test yourself**

7. Give an example of a viral disease.
8. Describe the symptoms of measles.
9. What does HIV stand for?
10. Describe the symptoms of the initial HIV infection.
11. Name a bacterial disease.
12. How is gonorrhoea treated?
13. Describe the symptoms of salmonella.
14. Describe the symptoms of gonorrhoea.

**Show you can...**

Explain how HIV affects those that catch it.
Fungal diseases

**Using a key to identify bacterial species**

Use the statement key below to identify the six species of bacteria.

1. It is made of more than one cell .................................. Go to question 2.
   - It is a single cell ................................................. Go to question 4.

2. It is made of two cells ................................ Streptococcus pneumoniae.
   - It is made of more than two cells .......................... Go to question 3.

3. Bacteria arranged in chains ................................ Streptococcus pyogenes.
   - Bacteria arranged in a cluster .............................. Staphylococcus aureus.

4. It has no flagella ................................ Treponema pallidum.
   - It has flagella ................................................. Go to question 5.

5. Its flagella are spread all around the cell ................ Salmonella typhi.
   - The flagella are located to one side of the cell .......... Helicobacter pylori.

**Figure 6.6 Bacterial species.**

---

**Fungal diseases**

Fungi are eukaryotic, like animals and plants, but unlike bacteria. They have evolved a huge range of appearances, from the single-celled yeast fungus to much larger multicellular mushrooms. Fungi can cause many diseases, including athlete’s foot in humans and rose black spot in roses.

**Rose black spot**

The fungus that causes this plant disease infects the leaves of the rose plant, causing them to develop purple or black spots which often turn yellow and fall off. Because the leaves are damaged, there is reduced photosynthesis and less growth. Fungal spores can travel by wind or water (e.g. rain splash) from plant to plant, causing the infection to spread. Rose black spot can be treated by using fungicides (chemicals that kill fungi) and removing and destroying infected leaves once they are first noticed.
Protist diseases

Protists are eukaryotic microorganisms. They are always unicellular or multicellular without tissues. This separates them from fungi, animals and plants. Protists are perhaps best known as pathogens for causing malaria in humans, but they also cause similar diseases in other animals and also infect plants.

**Malaria**

Malaria is a disease caused by Plasmodium protists. Approximately 200 million cases of malaria occur each year the world over. About half a million people then die from this disease each year. Its symptoms include repeated episodes of fever, and can eventually result in death. The Plasmodium (protist) pathogens are transmitted from one individual to another by mosquitoes. Because the mosquito transmits the protist (which causes malaria) we call the mosquito a vector. Mosquitos bite and suck blood from organisms infected by Plasmodium and then pass the pathogen from their saliva to the blood of all other organisms they bite.

Prevention is usually by avoiding being bitten. Mosquito nets and insect repellent sprays containing insecticides are often used. Mosquitos lay eggs in water that does not move, such as stagnant pools or puddles. To help stop the spread of malaria these can be filled in. No vaccination currently exists. Draining marshy areas can also destroy mosquito breeding sites.

**Test yourself**

15 Give an example of a fungal disease.
16 Explain how this fungal disease spreads.
17 Give an example of a disease caused by a protist.
18 Describe the symptoms of malaria.
19 Describe how malaria is transmitted.

**TIP**

Malaria is spread from person to person by mosquito bites. An excellent control measure is therefore to prevent mosquitoes breeding.
The first line of defence

The first line of defence is your body’s natural barriers to infection. These are not specific to the infecting pathogen, and so we describe them as **non-specific**.

Skin

Your skin is an amazing organ that almost completely covers any outer part of you that is prone to attack from pathogens. As well as this, it insulates you, helps you regulate your temperature and is involved in the way your senses provide you with information. When we get a cut, the blood clots and the wound quickly seals over to restore the protective barrier of the skin.

In vulnerable parts of our bodies, we can produce an **antimicrobial secretion** that prevents pathogens entering. For example, our tears contain antimicrobial agents to protect our eyes.

Hairs and cilia

Pathogens that are breathed in through your nose and mouth are often stopped before they reach your lungs. Your **nose** has hairs and produces mucus, which acts as a filter, stopping larger particles containing pathogens. You blow your nose or sniff and swallow, moving this mucus and pathogens either out of your body or into your stomach. Hairs are physical barriers against infection.

If pathogens pass your saliva or hairs in your nose, they are often stopped by the ciliated cells lining the inside of your **trachea** and **bronchi**, the tubes that reach down to your lungs. Ciliated cells possess **cilia**, which are tiny hair-like projections that protrude into the airway. In between the ciliated cells are cells that produce **mucus**, which they pump into the airway. Many pathogens and other particles that have been breathed in get stuck in this mucus. The ciliated cells beat (or waft) their cilia in a rhythmical pattern, which propels the mucus back up the airway to the back of the throat. Cilia are physical barriers against infection.

Stomach acid

Your stomach acid does not actually digest your food. It provides the correct pH for protease enzymes to start digesting protein, but it also has a crucial role in the first line of defence. It is hydrochloric acid and is strong enough to kill many bacterial pathogens that enter your body through your mouth or nose. It is a chemical barrier against infection.

The second line of defence

The second line of defence is your defence against pathogens that have entered your bloodstream or tissues.
If you are infected a second time with the same pathogen, a special type of lymphocyte will recognise its antigens and be able to produce larger numbers of antibodies more quickly. We call these white blood cells ‘memory’ lymphocytes, and we produce these after we have been infected by a ‘new’ pathogen. This is likely to mean you won’t fall ill from the same pathogen twice. How do we get colds every winter, then? This is because there are several hundred different strains of common cold viruses that all have different antigens. You are not likely to have been infected by the same one, but lots of different ones that have similar symptoms.

Phagocytes

Phagocytes are another a type of white blood cell. They take in or engulf pathogens, as well as your own dead or dying cells. Phagocytes are attracted to any area of your body in which an infection is present. Antibodies cause pathogens to clump together. When a phagocyte comes into contact with a pathogen it binds to it. The membrane of the phagocyte then surrounds the pathogen and absorbs it into a vacuole within its cytoplasm. Enzymes are added to the vacuole to break down the pathogen. This process in which phagocytes break down pathogens clumped together by antibodies is called phagocytosis.
Antitoxins

Many pathogens produce toxins that also make you ill (as well as the pathogen itself). To defend against these specific toxins, your lymphocytes can produce a special type of antibody called an antitoxin. This will bind with and neutralise the toxin helping you feel better.

KEY TERM

**Antitoxin** A protein produced by your body to neutralise harmful toxins produced by pathogens.

Test yourself

20 Give an example of the first line of defence.
21 Name the two types of white blood cell.
22 Describe how stomach acid stops infection.
23 Describe the difference between chemical and physical barriers.

Vaccination

**KEY TERM**

**Vaccine** A medicine containing an antigen from a pathogen that triggers a low level immune response so that subsequent infection is dealt with more effectively by the body’s own immune system.

**TIPS**

- It is important that you can explain the use of vaccinations to prevent disease. You do not need to know specific times or dates for getting vaccines, though, or their side effects.
- The term ‘immunisation’ means ‘becoming immune to a pathogen’. This can be caused by vaccination but also naturally if you become infected by a pathogen and subsequently recover.

It is likely that you will have been vaccinated against a number of diseases since you were born. Typically you may have had the measles, mumps and rubella (MMR) vaccine at about 12 months old, and a second combined vaccine for diphtheria, tetanus, whooping cough and polio at about 3 years. These are all life-threatening diseases, and so vaccination is important.

If the vast majority of people in a population have a vaccination, then even if a small number of people become infected the disease is not likely to spread. This is called **herd immunity**. The reverse is also true. If few people have a vaccine and a small number become infected the disease will spread much more quickly.

A vaccine is a small quantity of a dead, inactive or genetically modified version of a pathogen. Crucially it must have the same antigens as the pathogen, or your body would not recognise the pathogen later. You are injected with this and your immune response begins. Lymphocytes produce antibodies and antitoxins. Because this process takes several days, you may feel slightly unwell after a vaccination.
However, if the same pathogen were to get past your first line of defence and infect you in the future, your 'memory' lymphocytes would respond by producing antibodies and antitoxins fast enough for you not to fall sick. This is your secondary immune response. Thus, vaccinations prepare your immune system in case you are infected later in life.

For many vaccines, you are likely to have had booster injections several years after the first injections. These serve as a timely reminder for your immune system and 'refresh' your memory lymphocytes.

▲ Figure 6.13 This graph shows the rate at which antibodies are produced after the first and second exposures to a pathogen.

Test yourself
24 What does ‘MMR’ in the MMR vaccine stand for?
25 Which cells produce antibodies?
26 Define the term ‘vaccine’.
27 Describe the difference between antigens and antibodies.

Antibiotics and painkillers

A drug is any substance that has a biological effect on the organism taking it. We do not normally include foods in this definition, even though many do affect people. Some drugs are natural, such as nicotine in tobacco. Others are manufactured, such as Viagra. Some have a positive effect on the taker, such as medicines, while others have negative side effects, such as cocaine. Some drugs are recreational. They are taken to alter a person’s mood or emotions. Legal recreational drugs include caffeine and nicotine. Illegal recreational drugs include cannabis and cocaine.

Antibiotics

Antibiotics are a very important group of medicines which kill bacteria. They have no effect on viruses and so should not be prescribed for the common cold or other viral diseases.
Different antibiotics attack bacteria in different ways. Penicillin (see below) is a commonly used antibiotic. It makes the cell walls of the bacteria weaker, and so they burst and are killed. Other antibiotics alter bacterial enzymes, and others stop bacteria from reproducing.

**Penicillin and Alexander Fleming**

Sir Alexander Fleming first discovered penicillin in 1928. He won the Nobel Prize for this discovery and penicillin has saved countless numbers of lives ever since. Fleming returned to his laboratory where he was studying bacterial growth. Legend has it that one of his Petri dishes was mistakenly left open and had accidently been contaminated by the fungus *Penicillium notatum*. Where the fungus grew the bacteria did not. Instead of simply throwing the Petri dish away, Fleming realised that the fungus was naturally producing a chemical that killed bacteria. This chemical was eventually refined to become the first antibiotic drug, penicillin.

**Antibiotic resistance**

Since Fleming’s discovery scientists have developed a large number of other antibiotics. These have saved many, many lives in recent years and led to the near removal of some major diseases such as tuberculosis. However, until very recently, we had not discovered any new antibiotics in 30 years. During this time some pathogens have been evolving to be resistant to our antibiotics. **Antibiotic-resistant bacteria** are difficult to treat, which is a major worry for doctors and scientists.

Drug companies are working extremely hard to find new antibiotics or alternatives to them.

Although antibiotics are used to kill bacteria, the fact that they do not kill viruses makes it hard to treat viral diseases. Also, because pathogenic viruses live in body cells and tissues, it is difficult to kill the viruses without damaging the cells.

---

**KEY TERM**

**Antibiotic-resistant bacterium**

A bacterium that cannot be killed by antibiotics.
Painkillers are drugs that relieve pain but do not kill the pathogens. Some painkillers are naturally found in plants. Aspirin comes from the bark of the willow tree. The naturally occurring compound is called salicin and was first discovered in 1763. It was not manufactured as aspirin until 1897. It is now one of the most widely used medicines in the world, with over 40,000 tonnes consumed each year. As well as relieving pain, it can reduce fever, swelling and inflammation. It is also used as a preventative drug for reducing the risk of heart attacks.

Other painkillers have been manufactured. A second common painkiller is paracetamol. Like aspirin, this is a mild painkiller often used to stop headaches or minor pain in other parts of the body. It is a major ingredient in many cold and flu remedies. Unfortunately, it is easy to take too much paracetamol at a time, which can cause fatal liver damage. This is why it is very important not to take more than the stated dose.

Both aspirin and paracetamol are ‘over-the-counter’ medicines, which means you can buy them in small numbers in chemists and supermarkets. Other stronger painkillers exist, such as tramadol and morphine. The use of these can only be prescribed by a doctor or used in accident and emergency situations.
Traditionally, drugs were extracted from plants and microorganisms, for example the heart drug digitalis originates from foxgloves. Digitalis can be used to treat people with irregular heartbeats. Now scientists in the pharmaceutical industry make new drugs, although the starting point is often a plant compound.

Modern drug development

Drug development is the process of identifying a new drug, testing it and then manufacturing it for sale. This is a very costly and time-consuming process. To get a drug to the stage at which it can be tested, the development is likely to take many years and hundreds of millions of pounds. Drug companies are some of the largest companies in the world. Many potential drugs don’t pass though the stages described below. From perhaps ten thousand possible drugs only 10 (0.1%) may ever get to be tested on humans.

The first stage of drug development involves computer modelling. The structure of the drug and the interactions it might have on naturally occurring substances in the body are looked at on a computer.

This initial stage of drug development also involves testing in the laboratory. This may include testing on live cells taken from an organism and grown in a Petri dish. Once it appears that the drug may be able to work on live cells in a laboratory, it is then tested on animals. This is a very important stage, because it allows the scientists to test the drug on a living organism without risking testing it on humans at this stage. These first stages of drug testing are called preclinical testing.

The final stage involves clinical trials on humans. There are three phases here. In the first, small amounts of the drug are tested on a small number of healthy volunteers to determine whether it is toxic and has any side effects, and to identify safe dosing volumes. In the second phase the drug is given to small numbers of sick patients to test how well it works (its efficacy). Finally, in phase three trials it is given to a large number of patients to finalise safe doses and efficacy. If a drug passes all of these stages it will be given a licence and then can be manufactured and sold.

Some of these trials are described as double blind. This is because the patients are randomly allocated to receive either the drug or a placebo (which looks like the drug but does not contain it) and the doctors also don’t know which patients are receiving which. This means those patients who receive the placebo are in effect a control group.

Test yourself

31 What is the first stage of drug development?

Show you can...

Describe how drug development occurs.
Chapter review questions

1. Define the term ‘communicable disease’.
2. Name the four types of pathogen that can cause communicable diseases.
3. Give an example of a disease caused by a fungus.
4. Give an example of a disease caused by a protist.
5. What does HIV/AIDS stand for?
6. Name the scientist who first discovered antibiotics.
7. Name the microorganisms upon which antibiotics don’t work.
8. Explain why medicines are often expensive.
9. How is salmonella often transmitted?
10. What are the symptoms of gonorrhoea?
11. What are the symptoms of malaria?
12. Describe how the transmission of malaria is reduced.
13. Describe how stomach acid acts in the first line of defence.
14. Explain how your immune system helps you ‘remember’ your previous infections.
15. Describe what a vaccine is.
16. Explain the idea of herd immunity.
17. Explain the purpose of booster injections.
18. Describe how antibiotics were first discovered.
19. Describe how cilia and mucus-producing cells prevent infection.
20. Describe how phagocytes prevent infection.
21. Describe how lymphocytes prevent infection.
22. What is the first step in drug development?
23. Define the term ‘efficacy’.
24. Explain what a double blind trial is.
Practice questions

1 Figure 6.19 shows data from a drug trial involving a new asthma medicine called Breathrite.

![Figure 6.19](image)

### a) Before the drug was trialled using people it was tested in the laboratory. What is the purpose of these tests? [1 mark]

### b) The drug trial used a double blind test. Doctors tested the effectiveness of Breathrite against a placebo using two groups of volunteers over 12 weeks.

1. What is a placebo? [1 mark]
2. Why is a placebo used? [1 mark]
3. Describe what a double blind trial means. [1 mark]

### c) It is important that the two groups of volunteers are similar. Give one factor that should be similar in both groups. [1 mark]

### d) Over the 12-week period, how much did the lung capacity of the volunteers increase in those who were given Breathrite? [1 mark]

2 In 2014 the largest Ebola epidemic in history broke out and affected a number of countries in West Africa. Ebola is caused by the *Ebola* virus.

![Figure 6.20](image)

### a) Give the term given to organisms like *Ebola* that cause disease. [1 mark]

### b) Describe how viruses like *Ebola* cause illness. [1 mark]

### c) Currently there are no cures for Ebola.

1. Why don’t antibiotics get rid of Ebola? [1 mark]
2. Some of the symptoms of Ebola are headaches and muscle pain. Suggest what might be given to someone suffering from Ebola to relieve the symptoms. [1 mark]

### d) Scientists are trying to create a vaccine for Ebola by using an inactive form of the *Ebola* virus. Explain how this would allow a person to become immune to the disease. [3 marks]

3 a) Malaria is a disease cause by what type of pathogen?

   A Virus  B Bacteria  C Fungi  D Protist [1 mark]

b) Which two of the following are symptoms of this disease?

   A Fever  B Vomiting  C Flaking skin  D Swollen feet [1 mark]

4 The influenza virus is a microorganism that can cause the flu. The flu is an infectious disease, which can be spread quickly from person to person.

### a) i) Name two other microorganisms that can cause disease. [2 marks]

### ii) Describe two ways in which microorganisms like influenza can be passed on from one person to another. [2 marks]

### iii) Suggest a simple hygiene measure that can be taken to reduce the spread of the flu. [1 mark]

### b) The body has several non-specific ways of preventing infection from microorganisms. Copy and complete Table 6.1 by naming the part of the body described. [3 marks]

<table>
<thead>
<tr>
<th>How entry of microorganisms is prevented</th>
<th>Part of the body</th>
</tr>
</thead>
<tbody>
<tr>
<td>Contains acid to destroy microorganisms</td>
<td></td>
</tr>
<tr>
<td>Acts as a barrier</td>
<td></td>
</tr>
<tr>
<td>Contain ciliated cells to trap bacteria</td>
<td></td>
</tr>
</tbody>
</table>

### c) If microorganisms make it past our body’s defences, specific cells are involved in protecting us from harm.

1. What is the name given to these types of cell? [1 mark]

   i) Describe the ways in which these cells can protect us from infectious disease. [3 marks]
Evaluating the risks and benefits

The World Health Organization (WHO) has the goal of eliminating measles in WHO regions between 2015 and 2020. Measles is a highly infectious disease caused by the measles virus and is spread through droplet inhalation and contact with infected people and surfaces. Measles is not just a condition that causes spots; about one in five children infected with it experience complications and one in ten can end up in hospital. In rare cases measles can cause death. Anyone can catch measles, and there is no specific treatment for it. The most effective way of preventing it is to have two doses of the combined MMR (measles, mumps and rubella) vaccination, which gives almost total protection from the disease.

In recent years there has been a global decline in MMR vaccinations, especially in Western Europe and the USA. This has contributed to over 22,000 cases of measles worldwide and has raised concerns that, far from being eliminated, cases of measles are actually increasing. According to WHO, a growing number of parents are refusing to vaccinate their children. This is sparking a global debate over whether vaccinations should be compulsory.

So what do you think?

In making any decision a number of factors need to be examined. These can be split into advantages, disadvantages and risks. These need to be considered on a personal level for the person making the decision, as well as in terms of the impact on society as a whole.

Questions

1 Copy the table below and use the internet and other sources of evidence, such as medical leaflets and newspaper articles, to help you determine the advantages, disadvantages and risks of compulsory vaccination programmes.

<table>
<thead>
<tr>
<th>Vaccination</th>
<th>Advantages</th>
<th>Disadvantages</th>
<th>Risks</th>
</tr>
</thead>
<tbody>
<tr>
<td>Compulsory</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Voluntary</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

2 Make a decision on what you think and write a letter to the Government expressing your views on compulsory vaccinations. Ensure you support your views with reasons and evidence.

3 Using diagrams and your knowledge of the immune system, explain how the MMR vaccination creates immunity to measles.
Imagine being able to make your own food. Not make it from ingredients in your kitchen, but actually make the ingredients themselves. This is what plants and other photosynthesising species can do. In fact, almost all life on Earth depends upon their ability to do this. They feed themselves by storing the Sun’s light as glucose during photosynthesis. Interestingly we could not live here without them, but they would grow equally well without us. In fact they might grow a little better!

This chapter covers specific points 4.4.1.1 to 4.4.1.3 and is called Photosynthesis. It covers photosynthesis, its rate of reaction and the uses of the glucose that it produces.
Photosynthesis is a chemical reaction that occurs in the green chloroplasts of plants and algae. It needs light from the Sun and converts the reactants carbon dioxide and water into glucose and the by-product oxygen. The word equation for photosynthesis is:

\[
\text{light} \quad \text{carbon dioxide + water} \quad \xrightarrow{\text{photosynthesis}} \quad \text{glucose + oxygen}
\]

The fully balanced symbol equation for photosynthesis is:

\[
\text{light} \quad 6\text{CO}_2 + 6\text{H}_2\text{O} \quad \xrightarrow{\text{photosynthesis}} \quad \text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2
\]

**Reactants and products**

Because it requires light and the green chlorophyll pigment in chloroplasts, photosynthesis mainly occurs in the leaves of plants. The palisade mesophyll cells (the cells near the top of leaves) have the highest number of chloroplasts and so most photosynthesis occurs here. Gases, including carbon dioxide, are absorbed from the air and enter the leaves via the stomata, before diffusing through the air spaces to reach the palisade cells. Water is absorbed by osmosis into the root hair cells before being transported to the leaves (and rest of the plant) through xylem vessels in transpiration.

Plants complete photosynthesis to produce glucose. They require light to do this, so photosynthesis cannot occur in the dark. Because it requires energy to work, it is an **endothermic reaction**. These are
TIP
Plants produce oxygen as a by-product. They don’t produce it to help animals live. It is handy for animals that they do, however, for without them there would be no oxygen to sustain their life on Earth.

Show you can...
Describe how oxygen levels would fluctuate during a 24-hour period.

Rate of photosynthesis

A limiting factor is anything that reduces the rate of a reaction. If you were making cakes but ran out of eggs, you couldn’t make any more cakes. So eggs would be the limiting factor. There are four possible limiting factors in photosynthesis:

1. low temperatures
2. shortage of carbon dioxide
3. shortage of light
4. shortage of chlorophyll.

At lower temperatures all chemical reactions occur more slowly because the reactant molecules have less kinetic energy, so they collide less, and therefore react less. Carbon dioxide is a reactant and so the reaction occurs slowly if carbon dioxide is in short supply. Light provides the energy necessary for this reaction so reduced levels mean reduced photosynthesis, and no light means no reaction. Plants absorb magnesium by active transport from the soil. They use this to make chlorophyll, so plants with a magnesium deficiency cannot photosynthesise as rapidly as those that have plenty of magnesium.

KEY TERM
Limiting factor Anything that reduces or stops the rate of a reaction.

KEY TERM
Deficiency A lack or shortage.

Test yourself
1. Name the chemical pigment required for photosynthesis to occur.
2. Give the chemical formulae of the products of photosynthesis.
3. Explain why photosynthesis cannot happen in the dark.

Diagram:

- Figure 7.2 (a) Effect of temperature on the rate of photosynthesis at a fixed light intensity and fixed carbon dioxide concentration.
- (b) Effect of increasing light intensity on the rate of photosynthesis at a fixed temperature and fixed carbon dioxide concentration.
Food production is an incredibly important job in all countries of the world. Farmers and growers of crop plants will do as much as they can to maximise the useful part of their crop (the yield) or make their plants grow as quickly as possible. To do this they ensure that plant growth is not limited by a limiting factor. They maintain their plants in greenhouses or polytunnels, which allow them to:

- be grown at the optimum temperature (using heaters if necessary)
- have burners near their plants to produce extra carbon dioxide for optimal growth
- provide their plants with maximum light (using artificial lighting if necessary).

Farmers and growers must balance these three conditions without making their crops too expensive to sell at a profit.

The four limiting factors described on the previous page often interact. Any of them may, individually or together with another factor, be responsible for limiting the rate of photosynthesis.

![Figure 7.3](image_url) How the rate of photosynthesis is affected by increasing light intensity: (a) on a hot and a cold day; (b) at higher and lower CO₂ levels.

**KEY TERM**

**Yield** The amount of an agricultural product.

**Figure 7.4** Farmers ensure their plants have optimum conditions for maximum growth.

**Test yourself**

4 Define the term ‘limiting factor’.
5 Name the limiting factors for photosynthesis.
6 Describe why plants need magnesium.
7 Describe how farmers can increase the rate of photosynthesis of plants inside a glasshouse.
Investigate the effect of light intensity on the rate of photosynthesis using an aquatic organism such as pondweed

In this practical you will investigate the effect of light intensity on the rate of photosynthesis.

Method
1. Take a piece of pondweed (Cabomba or Elodea) and cut it underwater so that the stalk is cut at a 45° angle.
2. Keeping it under the water, attach a paper clip at the opposite end to the cut stalk and transfer to a boiling tube.
3. Set up a lamp, metre ruler and tank or beaker of water as shown in Figure 7.5.
4. Position the boiling tube with pondweed in it so it is 50 cm away from the light source.
5. Turn off the classroom lights, turn the lamp on and wait for 2 minutes.
6. Ensure that the temperature of the water in the boiling tube with the pondweed remains constant throughout the experiment by monitoring with a thermometer.
7. Count the number of bubbles produced in 1 minute. Repeat this twice more so that a mean rate of bubbles produced can be determined.
8. Move the clamped boiling tube so that it is 40 cm away from the light source and repeat steps 5 and 6.
9. Repeat the experiment for a distance of 30, 20 and 10 cm.
10. Record your results in a table and draw a graph of mean number of bubbles against distance of lamp.

Questions
1. Use the equation $\frac{1}{\text{distance}^2}$ to determine the light intensity for each distance used. For example, if the distance were 60 cm the light intensity would be $\frac{1}{60^2}$ or $\frac{1}{(60 \times 60)}$.
2. Plot a line graph of your data with light intensity on the x-axis and mean rate of bubbles on the y-axis.
3. What can you conclude from your data about how light intensity affects photosynthesis?
4. Why was a beaker of water placed between the lamp and the pondweed?
5. Why did you have to wait 2 minutes before counting the number of bubbles?
6. Suggest how you could modify this experiment to measure the rate of oxygen production more accurately.
Uses of glucose from photosynthesis

It often helps to think of photosynthesis in terms of where the energy is. The two reactants, carbon dioxide and water, have a relatively small amount of energy within them. The energy found within the two products, glucose and oxygen, is greater. You will have learnt that energy cannot ever be created or destroyed, only converted from one form to another. So how can the products of this reaction have more energy than the reactants? Where does this energy come from?

The answer is energy transferred by the light. This energy is used to power the reaction. It breaks apart the chemical bonds in the reactants and re-forms them with some extra energy stored in the bonds of the products. It is this extra energy that supports almost all life on Earth. When you think like this, it is obvious that photosynthesis can’t happen in the dark.

The energy that plants have stored in the formation of glucose during photosynthesis has five general uses. It is:

1. used in respiration
2. converted into insoluble starch and stored
3. converted into fats and oils and stored
4. used to make cellulose which strengthens the cell walls in plants
5. used with nitrate ions absorbed from the soil to make amino acids for protein synthesis.

Show you can...

State whether there is more energy in the reactants or products of photosynthesis.
Chapter review questions

1. Name the two reactants in photosynthesis.
2. Name the plant organs in which most photosynthesis occurs.
3. Name the plant cell organelle in which photosynthesis occurs.
4. What is the source of the energy required for photosynthesis to occur?
5. Explain why there would be almost no life on Earth without plants.
6. Give the word equation for photosynthesis.
7. Name the green chemical that must be present for photosynthesis to occur.
8. Name the plant tissue in which most photosynthesis occurs.
9. Name the plant tissue with the highest concentration of chloroplasts.
10. Name the type of specialised cell in a plant that absorbs the most water.
11. Give the chemical formula for glucose.
12. Define the term ‘limiting factor’.
13. Name the process by which plants absorb mineral ions from the soil.
14. Name the key metal element in making chlorophyll.
15. Explain why farmers often grow crops in polytunnels and greenhouses.
16. Give the five uses of glucose formed during photosynthesis.
17. Name the compound produced by plants to strengthen their cell walls.
18. Describe why photosynthesis is an endothermic reaction.
19. Give the balanced symbol equation for photosynthesis.
20. Name the four limiting factors in photosynthesis.
21. Explain why less photosynthesis occurs at lower temperatures.
22. Explain why more photosynthesis occurs with more carbon dioxide.
23. Explain why less photosynthesis occurs under lower light conditions.
24. Describe an experiment in which you investigate the effect of light intensity on photosynthesis.
25. Define the term ‘yield’.
26. Describe what farmers can do to greenhouses or polytunnels in which they are growing plants to improve their yield.
27. Explain how we know there is more energy in glucose and oxygen combined than in carbon dioxide and water combined.
28. Name the type of ions absorbed by roots which are used by plants to make proteins.
Practice questions

1  a) Copy and complete the word equation for photosynthesis. [2 marks]
   carbon dioxide + X → glucose + Y
 b) Describe how the carbon dioxide needed for photosynthesis gets into the plant. [1 mark]
 c) Copy and complete the sentences by choosing the correct words from the box.

light root mitochondria respiration chemicals flower chloroplasts haemoglobin chlorophyll ribosome leaf

   i) The plant organ that is specialised to carry out photosynthesis is the ___________________________. [1 mark]
   ii) The energy needed for photosynthesis to occur comes from ___________________________. [1 mark]
   iii) Energy is absorbed by a green pigment called ___________________________. [1 mark]
   iv) This green pigment is found in small organelles called ___________________________. [1 mark]

2  After a plant had been kept in the dark for 48 hours it was set up as shown in Figure 7.7. After another 24 hours the leaves from inside bags A, B and C were removed and a test was carried out to see if starch was present.

   a) Copy and complete Table 7.1 to show the likely results of the experiment by placing the letters in the correct column. [2 marks]

<table>
<thead>
<tr>
<th>Starch present</th>
<th>Starch not present</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

   b) i) Which chemical would have been used to test the leaves for starch? [1 mark]
        A  Biuret solution   C  Iodine solution
        B  Benedict’s solution D  Ethanol solution
   ii) What colour would indicate a positive result for starch? [1 mark]
        A  Purple            C  Reddish orange
        B  Blue-black         D  Green

3  A farmer was growing tomatoes in a greenhouse. The graph shows the effect of temperature and concentration of carbon dioxide on the rate of photosynthesis of the tomato plants.

   a) From the graph conclude:
      i) the best temperature at which to grow the tomato plants [1 mark]
      ii) the best concentration of carbon dioxide to use to grow the tomato plants [1 mark]
      iii) the maximum rate of photosynthesis recorded. [1 mark]
   b) Apart from temperature and carbon dioxide concentration, what other factor could the farmer change to increase the rate of photosynthesis? [1 mark]

4  Outline a plan to investigate how the rate of photosynthesis in pondweed changed when the intensity of light was changed.

   a) Describe how you would do the investigation and the measurements you would take. [6 marks]
   b) Describe how you would make it a fair test.
Planning and variables

A variable is a quantity or characteristic. In investigations three types of variable are discussed: **independent variables**, **dependent variables** and **control variables**.

By knowing what the variables are, a clear method can be written that outlines the procedure and the steps taken in order to make it a fair test.

A group of students were planning to investigate the effect of carbon dioxide concentration on the rate of photosynthesis. They decided they would measure the rate of photosynthesis by hole-punching small discs out of spinach leaves, and placing them in a 10 cm³ syringe of sodium hydrogen carbonate solution (NaHCO₃). By pressing on the plunger of the syringe they could increase the pressure. This forced any trapped air out of the spongy mesophyll of the leaf discs and made them sink. Because gas is produced in photosynthesis, the time taken for the leaf discs to rise (inflate again) could be used as a measure of the rate of photosynthesis.

Before they started, the students wrote up their experiment method.

1. Read through the method below and write down any parts where it is not clear how a control variable will be kept constant. Explain how the method could be improved to overcome this.
2. Then write down a clear method to explain how the above experiment could be modified to investigate the effect of temperature on the rate of photosynthesis. Ensure you clearly detail what the variables would be.

**Method**

1. Wearing eye protection, cut out the spinach leaves using a hole-punch and add these to the syringe by removing the plunger.
2. Add to this one drop of washing-up liquid and replace the plunger.
3. Draw up a 0.5% sodium hydrogen carbonate solution into the syringe.
4. Hold the syringe with the tip pointing upward and expel any air that remains in the syringe by depressing the plunger carefully. Stop before any of the sodium hydrogen carbonate solution comes out.
5. Place a finger over the tip of the syringe and hold firmly in place. Pull back on the plunger and hold for 10–15 seconds to create a partial vacuum inside the syringe. Look to see signs of bubbles escaping the edges of the leaf discs.
6. Release your finger and the plunger at the same time, tap gently on the side of the tube and the leaf discs should start to sink.
7. Repeat steps 5 and 6 until all the leaf discs have sunk.
8. Place the syringe under a lamp and time how long it takes for the first leaf disc to reach the surface.
9. Repeat the experiment using a 0.4, 0.3, 0.2 and 0.1% concentration of sodium hydrogen carbonate solution.

**Questions**

1. What is the dependent variable in this experiment?
2. What is the independent variable in this experiment?
3. What variables must be controlled in order to make the experiment a fair test?
Every one of the thousands of billions of cells that make up your body is respiring now and will continue to do so. If a cell stops it will die. Respiration is therefore an extremely important reaction. It is one of the seven life processes. Just like you, all other life on our planet undergoes respiration or a chemically similar reaction to release the energy it needs to survive. When we discover life elsewhere in the Universe it will probably do something similar.

This chapter covers specification points 4.4.2.1 to 4.4.2.3 and is called Respiration.

It covers aerobic and anaerobic respiration, response to exercise and metabolism.
Aerobic respiration

Respiration is a chemical reaction that occurs in the mitochondria of your cells. This reaction releases the energy stored in glucose to allow your cells to complete chemical reactions and to allow animals to move and keep warm. The word equation for aerobic respiration is:

\[ \text{glucose} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water} \]

The fully balanced symbol equation for aerobic respiration is:

\[ C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O \]

Respiration is an oxidation reaction, because it uses oxygen.

Reactants and products

Plants and algae photosynthesise to store energy in glucose. Respiration releases this energy by reversing the process. Glucose reacts in the presence of oxygen to form carbon dioxide and water.

KEY TERMS

Aerobic In the presence of oxygen.
Oxidation A reaction that uses oxygen.

TIPS

- Some people confuse the process of breathing, called ventilation, with the release of energy from glucose in the chemical reaction called respiration.
- Respiration occurs continuously in all living cells.

Prior knowledge

Previously you could have learnt:

- Aerobic and anaerobic respiration are important processes in living organisms, that involve the breakdown of organic molecules produce energy.
- Aerobic respiration is an important process in living organisms that can be described by a word equation.
- Anaerobic respiration occurs in humans and microorganisms (e.g. fermentation), and can be described by a word equation.
- Aerobic and anaerobic respiration have differences in terms of reactants used, products formed, and implications for the organism.

Test yourself on prior knowledge

1. Name the products of aerobic respiration.
2. Describe the difference between aerobic and anaerobic respiration.
3. Explain how we use organisms that respire anaerobically. Give an example in your answer.
Unlike photosynthesis, which only occurs in the light, your cells must respire continuously through the day and night. Respiration does not require energy to drive the reaction like photosynthesis does. In fact, respiration releases energy to allow respiring cells to live. This makes it an exothermic reaction.

It is easy to remember the reactants by answering these questions:
1 Why do I eat? Answer: To get glucose (and other key nutrients).
2 Why do I breathe? Answer: To get oxygen.

It is easy to remember the products by picturing yourself breathing out on to a cold window. What two substances do you breathe out? Carbon dioxide and water (vapour which condenses on the window).

**TIP**
Aerobic respiration does not mean ‘in the presence of air’, even though it sounds like it does. It means ‘in the presence of oxygen’.

Exo sounds like exit. You leave through an exit, so energy leaves in an exothermic reaction.

**KEY TERM**
Exothermic reaction A reaction that gives out thermal energy.

**Test yourself**
1 What type of reaction is respiration?
2 Name the type of cellular organelle in which respiration occurs.
3 Define the term ‘aerobic’.
4 State the balanced symbol equation for respiration.

**Investigating respiration in invertebrates**
Students used a simple respirometer to investigate how much oxygen is used by different invertebrate species. This is shown in Figure 8.2.

**Questions**
1 a) Which gas is produced by both invertebrates during the experiment?
   b) What happens to this gas during the experiment?
2 a) What happens to the water drops in the capillary tubes over the course of the experiment in both respirometers?
   b) Explain the reason for this movement.
3 Which respirometer gives the more accurate reading? Explain why.
   Note: the grasshopper and cricket were released from the respirometers before they ran out of oxygen, and no insects were harmed during this experiment!

**Figure 8.1** Water is produced during respiration and condenses on windows when you exhale.

**Figure 8.2** The equipment used to investigate how much oxygen is used by different invertebrate species.
Conversion of energy in respiration (and photosynthesis)

So the reactants and products in respiration are the opposite of those in photosynthesis.

Photosynthesis:

```
light

carbon dioxide + water → glucose + oxygen
```

Respiration:

```
energy

glucose + oxygen → carbon dioxide + water
```

However, these two crucial chemical reactions are not simply the opposite of each other when we focus on the flow of energy. Photosynthesis is an endothermic reaction, which requires energy from its surroundings to occur. The arrow on the photosynthesis equation shows energy in. Respiration is an exothermic reaction, which releases energy. The arrow on the respiration equation shows energy out.

These two equations work beautifully together for us and many other organisms on Earth. They show how:
- energy transferred by light (mainly from the Sun) is converted into a chemical store of energy in glucose by photosynthesis
- energy for life processes is released from glucose by respiration.

The energy released from glucose has two main functions. It is converted into:
- thermal energy to keep an organism warm (especially in warm-blooded birds and mammals)
- a chemical store of energy that is available for reactions and processes, such as movement, in an organism.

Investigating the temperature rise caused by respiration in yeast

To investigate the temperature rise produced by respiring organisms, a student heated 200 cm$^3$ of a 10% glucose solution to 35 °C and then stirred in 20 g of baker’s yeast (Saccharomyces cerevisiae). The mixture was then poured into a thermos flask and a tightly fitting bung with two bore holes was used to stopper the flask. A thermometer was inserted in one of the holes so the temperature could be monitored over the course of an hour.

Questions
1. Explain why a second hole was needed in the bung.
2. a) Describe how you would expect the temperature readings to vary over the course of an hour.
   b) Explain why this is.
3. Describe a suitable control that could be used in this experiment and explain the purpose of having one.
Anaerobic respiration

○ Response to exercise

When we exercise, we use up energy more quickly. To produce the extra energy, we respire more quickly and therefore use up oxygen and glucose faster. We **breathe faster** and **more deeply** to take in more oxygen. Our **hearts beat faster** to deliver glucose and oxygen more quickly to our rapidly respiring muscles. There will be times in your life when you cannot breathe quickly or deeply enough to supply all of your cells with all the oxygen they need to keep on respiring aerobically. This might be the last time you had to run the cross-country at school or when you were last out of breath. At this point your cells, particularly your muscles, start to run out of oxygen. They can only respire **anaerobically**. The word equation for this is:

\[
\text{energy (only 5%)} \quad \text{glucose} \rightarrow \text{lactic acid}
\]

○ Reactants and products

Cells respiring anaerobically are missing oxygen and so cannot make carbon dioxide and water. Instead they make an intermediary substance called **lactic acid**. A build-up of lactic acid in a muscle causes cramp. This causes muscle **fatigue** and stops the muscle contracting so efficiently.

Because the reaction has not been fully completed (because of the lack of oxygen) the total energy released from anaerobic respiration is much less than during aerobic respiration. Only about 5% of the energy (or \(\frac{1}{20}\)) is released. This means that when you are respiring anaerobically, your body is releasing less energy and producing a substance that hurts you (lactic acid). Is it trying to tell you something?!

○ Paying your oxygen debt

When you have finished exercising vigorously it is likely that you will sit down and relax. We say you have an **oxygen debt** to your body at this point. (You owe it oxygen.) For the next few minutes you will continue to **breathe deeply** and **quickly** to replenish the oxygen you have used up. You are repaying your oxygen debt. Your **pulse rate** will remain **high** to pump the newly oxygenated blood and glucose as quickly as possible to your muscles for more aerobic respiration.

After a few minutes your breathing will return to normal. Now you probably feel tired but not nearly as tired as you did a few minutes ago.
The lactic acid that builds up during anaerobic respiration diffuses from a high concentration in your muscle cells to a low concentration in your blood. By the time it reaches your liver it is at a high concentration in your blood and so diffuses again into the low concentration in your liver. Here it is converted back into glucose by an oxidation reaction. This then diffuses into the blood for use in either aerobic or anaerobic respiration.

When you convert the lactic acid to glucose in the liver, you are paying back your oxygen debt and making the energy available that was locked up in the lactic acid.

**Test yourself**

9. Give the percentage of energy released in anaerobic respiration compared to that released in aerobic respiration.

10. Where is lactic acid broken down?

11. Describe where the majority of the energy is stored in anaerobic respiration.

12. Define the term 'anaerobic'.

**Show you can...**

Explain what happens from when your muscle cells start to respire anaerobically during a run until you recover completely afterwards.

**Anaerobic respiration in plants and microorganisms**

It is a common misunderstanding that plants only photosynthesise and animals only respire. If plants only photosynthesised they would all be mass-producing glucose but not be able to use this energy to complete the life processes (and so die). Just like each and every one of your cells, every plant cell in every plant on our planet must respire or it will die. So plants photosynthesise during the day, and respire during the day and also the night. Just like animals, plants respire anaerobically when they do not have enough oxygen. This occurs in root cells in waterlogged soils and can cause roots, and eventually the whole plant, to die. Some microorganisms such as yeast, a single-celled fungus, also respire in this way. The equation for their anaerobic respiration is:
If this reaction happens in yeast cells it is called fermentation. This microorganism and this reaction are economically important in the manufacture of alcoholic drinks. The compound ethanol that is produced is commonly known as alcohol.

Yeast is also economically important in the manufacture of bread. In bread making, the carbon dioxide produced by the yeast causes the dough (mixture of flour and water) to rise. If the same reaction is occurring, why does our bread not taste alcoholic? Bread is baked in an oven and so the yeast is killed before it can make too much ethanol. Any that it does make is evaporated away by the warmth of the oven.

**Test yourself**

13 Give the word equation for fermentation.
14 Describe why bread isn't alcoholic even though yeast ferments in its making.

**Show you can...**

Explain why some organisms such as yeast produce alcohol.

**Metabolism**

Metabolism is the sum of all the chemical reactions that happen in a cell or in your body. These reactions include digestion of food, aerobic and anaerobic respiration which you have just learnt about, and protein synthesis. For plant cells, metabolism also includes photosynthesis. Most of these reactions are controlled by enzymes.

- **Breakdown and synthesis reactions**

Some metabolic reactions make more, smaller, often less complicated molecules. These are breakdown reactions.

Other reactions make larger, more complex molecules. These are synthesis reactions. To occur, these reactions need energy from respiration.
### Table 8.1 Breakdown reactions and synthesis reactions.

<table>
<thead>
<tr>
<th>Type of reaction</th>
<th>Comment</th>
</tr>
</thead>
<tbody>
<tr>
<td>Breakdown reactions</td>
<td>Synthesis reactions (building up)</td>
</tr>
<tr>
<td>Carbohydrates</td>
<td>Complex carbohydrates, e.g. starch, are broken down into simple sugars, e.g. glucose. Glucose is broken down in respiration in plants and animals to release energy.</td>
</tr>
<tr>
<td>Proteins</td>
<td>Proteins broken down into amino acids.</td>
</tr>
<tr>
<td>Lipids (fats)</td>
<td>Lipids are broken down to glycerol and fatty acids. Each lipid molecule breaks down to form one glycerol molecule and three fatty acid molecules.</td>
</tr>
<tr>
<td>Note</td>
<td>The above breakdown reactions take place in the gut in animals to produce smaller, soluble sub-units that can be absorbed into the bloodstream.</td>
</tr>
</tbody>
</table>
Chapter review questions

1. Explain the difference between respiration and ventilation.
2. Name the reactants in aerobic respiration.
3. Define the term ‘aerobic’.
4. Give the word equation for respiration.
5. Describe the conditions under which anaerobic respiration occurs.
6. Give the word equation for anaerobic respiration.
7. What are the effects of producing too much lactic acid?
8. Name the process by which oxygen moves into the cells for respiration.
9. Give the word equation for anaerobic respiration in yeast.
10. Explain the economic importance of anaerobic respiration in yeast.
11. Name the organelle in which respiration occurs.
12. What are the uses of the energy released during respiration?
13. State why respiration is an exothermic reaction.
14. Give the proportion of energy released in anaerobic respiration compared with aerobic respiration.
15. Define the term ‘oxygen debt’.
16. Explain why your heart rate and breathing rate increase during exercise.
17. Define the term ‘fermentation’.
18. Define the term ‘metabolism’.
19. Describe an experiment in which you investigate how respiring yeast raises the temperature of the substrate it is growing in.
20. Explain the flow of energy through photosynthesis and aerobic respiration.
21. Name the main source of energy for the vast majority of the reactions in your cells.
22. Where is the unreleased energy stored in anaerobic respiration?
23. What happens to the lactic acid produced during anaerobic respiration?
24. Define the term ‘oxidation’.
25. Describe an experiment in which you investigate the oxygen consumption for respiration in two invertebrate species.
Practice questions

1 Respiration occurs in living organisms.
   a) i) What is the purpose of respiration? [1 mark]
   ii) Using your knowledge of the word equation for aerobic respiration, identify which of the following are chemical products. [2 marks]
      A Carbon dioxide   B Energy
      C Glucose           D Water
   iii) Glucose is a reactant in respiration. Which of the following is the correct chemical formula for glucose? [1 mark]
      A C₁₂H₂₀O₁₂      B C₆H₁₂O₆
      C C₆H₁₂O₆
   b) Figure 8.8 shows how a respirometer was set up to investigate aerobic respiration in germinating peas. Apparatus A shows the starting point of the water in the respirometer.

   ▲ Figure 8.8
   i) Predict which tube, B or C, shows the direction of liquid movement in the capillary tube after 10 minutes. [1 mark]
   ii) Explain the role of the soda lime in this experiment. [2 marks]

2 A student was investigating the effect of exercise on an athlete’s body. She collected data on their heart rate before and after exercise. Table 8.2 shows the data collected.

<table>
<thead>
<tr>
<th></th>
<th>Resting</th>
<th>During exercise</th>
</tr>
</thead>
<tbody>
<tr>
<td>Heart rate in beats per minute</td>
<td>65</td>
<td>125</td>
</tr>
<tr>
<td>Volume of blood pumped out of the heart in each beat in cm³</td>
<td>90</td>
<td>145</td>
</tr>
<tr>
<td>Cardiac output in cm³ per minute</td>
<td>5850</td>
<td></td>
</tr>
</tbody>
</table>

a) By how much did the athlete’s heart rate increase during exercise? [1 mark]

3 Yeast cells can respire anaerobically. This means they can be grown in anaerobic conditions inside a fermenter.
   a) What does anaerobic mean? [1 mark]
   b) i) Which of the following is the correct word equation for anaerobic respiration in yeast? [1 mark]
       A Glucose → lactic acid + carbon dioxide
       B Glucose + oxygen → lactic acid + carbon dioxide
       C Glucose → ethanol + carbon dioxide
       D Glucose + oxygen → ethanol + carbon dioxide
   ii) Give one way anaerobic respiration in yeast is different to anaerobic respiration in human cells. [1 mark]

4 Figure 8.9 shows the changes that occur in a fermenter over the course of 24 hours. Describe and explain how ethanol production changed over the 24-hour period. [2 marks]

   ▲ Figure 8.9

5 A woman is taking part in an exercise class. Initially her body cells are carrying out aerobic respiration.
   a) Write a balanced equation for aerobic respiration. [2 marks]
   b) As the exercise becomes more vigorous, her cells switch to anaerobic respiration.
      i) Name the chemical produced in anaerobic respiration that is not produced in aerobic respiration. [1 mark]
      ii) Explain how the substance produced is broken down after the exercise has finished. [2 marks]

6 Describe the similarities and differences between aerobic and anaerobic respiration in humans. [6 marks]
Means and ranges

Two students were investigating the effect of temperature on respiration in yeast. Terry set up his experiment using the apparatus shown in Figure 8.10a and Afreen as shown in Figure 8.10b.

Both students added the same amount of glucose solution and yeast to their apparatus and covered it using 20 cm³ of liquid paraffin. They then placed their equipment in a water bath at 20 °C and left it for 5 minutes. After the 5-minute period, they measured the volume of gas produced in 10 minutes: Terry by counting the number of bubbles and Afreen by recording the movement of the gas syringe. They repeated their experiment five times and then repeated it over a range of temperatures, from 20 to 80 °C.

Figure 8.10 The equipment used by Terry (a) and Afreen (b) to investigate the growth of yeast.
Both of the students' data show variation in results. This is to be expected with repeated results. The range of the results is the highest and lowest value recorded for each set of data. The narrower the range, the closer the results are to each other and the more repeatable the data. For Terry's results at 20°C the range is 345–400.

### Questions

1. Calculate the range for Terry's other temperature results and determine which of his data sets has the most repeatable data.

The more repeatable the data, the more confident you can be that the true value of a measurement lies within the range of your data. To estimate a true value from a range you should calculate a mean. This is done by adding up the repeats and dividing the number by the total number of repeats taken. The result should be rounded to the same number of decimal places or one more than the raw data.

### Questions

2. Calculate the means for Terry's results at 35°C, 50°C, 65°C and 80°C.

The more repeats taken, the easier it is to spot anomalies. An anomaly is a value that is not in line with the other data and is therefore not likely to be caused by random variation.

### Questions

3. Identify any anomalous results in Afreen's data.

When data can be identified as being anomalous they should be left out of the ranges and means, as they can skew the data and make them less accurate.

### Questions

4. Work out the ranges and means for Afreen's data.
5. What is the trend shown in both students’ data? Can you explain this?
6. Whose equipment allowed them to gather more accurate data? Explain why.
7. Why in both experiments was the yeast and glucose solution left for 5 minutes before the measurements were started?
8. Why in both experiments was the yeast and glucose solution covered in liquid paraffin?
9. What gas were both students collecting, and how could they test it to prove what it is?
This chapter covers specification points 5.1.1.1 to 5.1.2.6 and is called Atomic structure and the periodic table.

It covers the structure of atoms, reactions of elements, the periodic table and mixtures.

Writing formulae and equations is covered separately in Chapter 14.

Until you reached GCSE, Chemistry was studied at the particle level. In order to take chemistry further, you now need to understand what is inside atoms. The elements in the periodic table are ordered by what is inside their atoms. An understanding of the periodic table allows you to explain and/or work out a lot of chemistry even if you have never studied it.
Structure of atoms

Protons, neutrons and electrons

Atoms are the smallest part of an element that can exist. Atoms are made up of smaller particles called protons, neutrons and electrons. The table below shows the relative mass and electric charge of these particles. The mass is given relative to the mass of a proton. Protons and neutrons have the same mass as each other while electrons are much lighter (Table 9.1).

<table>
<thead>
<tr>
<th></th>
<th>Proton</th>
<th>Neutron</th>
<th>Electron</th>
</tr>
</thead>
<tbody>
<tr>
<td>Relative mass</td>
<td>1</td>
<td>1</td>
<td>very small</td>
</tr>
<tr>
<td>Relative charge</td>
<td>+1</td>
<td>0</td>
<td>−1</td>
</tr>
</tbody>
</table>

The structure of atoms

Atoms are very small. Typical atoms have a radius of about 0.1 nm (0.000 000 000 1 m, that is \(1 \times 10^{-10}\) m). Atoms have a central nucleus which contains protons and neutrons (Figure 9.1). The nucleus is surrounded by electrons. The electrons move around the nucleus in energy levels, or shells.
The nucleus is tiny compared to the size of the atom as a whole. The radius of the nucleus is less than 1/10000th of that of the atom (1 × 10^{-14} m). This difference in size between a nucleus and an atom is equivalent to a pea placed in the middle of a football pitch (Figure 9.2).

The nucleus contains protons and neutrons. These are much heavier than electrons. This means that most of the mass of the atom is contained in the tiny nucleus in the middle.

Test yourself
1. Carbon atoms have a radius of 0.070 nm. Write this in standard form in the units of metres.
2. The radius of a hydrogen atom is 2.5 × 10^{-11} m. Write this in nanometres.
3. The radius of a chlorine atom is 1 × 10^{-10} m and the radius of a silicon atom is 0.060 nm. Which atom is bigger?
4. Sodium atoms have a radius of 0.180 nm. The nucleus of an atom is about 10000 times smaller. Estimate the radius of the nucleus of a sodium atom. Write your answer in both nanometres and metres.
5. A copper atom has a diameter of 0.256 nm. A copper wire has a diameter of 0.0440 cm.
   a) Write the diameter of the atom and the wire in metres.
   b) How many times wider is the copper wire than a copper atom? Give your answer to 3 significant figures.
6. A gold atom has a diameter of 0.270 nm. The largest gold bar in the world is 45.5 cm long. How many gold atoms fit into 45.5 cm? Give your answer to 3 significant figures.
Atomic number and mass number

The number of protons that an atom contains is called its atomic number. Atoms of different elements have different numbers of protons. It is the number of protons that determines which element an atom is. For example, all atoms with 6 protons are carbon atoms, while all atoms with 7 protons are nitrogen atoms.

All atoms are neutral, which means they have no overall electric charge. This is because the number of protons (which are positively charged) is the same as the number of electrons (which are negatively charged).

Most of the mass of an atom is due to the protons and neutrons. Protons and neutrons have the same mass as each other. The mass number of an atom is the sum of the number of protons and neutrons in an atom. For example, an atom of sodium has 11 protons and 12 neutrons and so has a mass number of 23.

\[
\text{ATOMIC NUMBER} = \text{number of protons} \\
\text{MASS NUMBER} = \text{number of protons} + \text{number of neutrons}
\]

The atomic number and mass number of an atom can be used to work out the number of protons, neutrons and electrons in an atom:

- number of protons = atomic number
- number of neutrons = mass number – atomic number
- number of electrons = atomic number (only for atoms, not ions).

The mass number and atomic number of atoms can be shown as in Figure 9.3.

Example

How many protons, neutrons and electrons are there in an atom of \(^{81}\)Br?

Answer

Number of protons: 35 (we could also find this by looking at the atomic number in the periodic table if it was not shown).

Number of neutrons: 81 – 35 = 46 (the mass number minus the number of protons).

Number of electrons: 35 (the same as the number of protons).

Isotopes

For most elements there are atoms with different numbers of neutrons. Atoms with the same number of protons but a different number of neutrons are called isotopes. This means that isotopes have the same atomic number but a different mass number.

For example, carbon has three isotopes and so there are three different types of carbon atoms. These are shown in the table below. These three isotopes are all carbon atoms because they all contain 6 protons, but they each have a different number of neutrons (Table 9.2).
Isotopes have a different mass and mass number, but their chemical properties are the same because they contain the same number of electrons.

Relative atomic mass

The relative atomic mass \((A_r)\) of an element is the average mass of atoms of that element taking into account the mass and amount of each isotope it contains. This can be calculated as shown:

\[
A_r = \frac{\text{total mass of all atoms of element}}{\text{total number of atoms of that element}}
\]

**Example**

Find the relative atomic mass of chlorine which is found to contain 75% of atoms with mass number 35, and 25% of atoms with mass number 37. Give the answer to one decimal place.

**Answer**

\[
A_r = \frac{[75 \times 35] + [25 \times 37]}{75 + 25} = \frac{3550}{100} = 35.5
\]


**Electron arrangement**

The **electrons** in an atom are in **energy levels**, also known as **shells**. Electrons occupy the lowest available energy levels. The lowest energy level (the first shell) is the one closest to the nucleus and can hold up to two electrons. Up to eight electrons occupy the second energy level (the second shell) with the next eight electrons occupying the third energy level (third shell). If there are two more electrons they occupy the fourth energy level.

The arrangement of electrons in some atoms are shown in Table 9.4. The **electronic structure** can be drawn on a diagram or written using numbers. For example, the electronic structure of aluminium is 2,8,3 which means that it has two electrons in the first energy level, eight electrons in the second energy level and three electrons in the third energy level (Table 9.4).

Table 9.4 The electronic structures of some atoms.

<table>
<thead>
<tr>
<th>Atom</th>
<th>Atomic number</th>
<th>Number of electrons</th>
<th>Electronic structure (written)</th>
<th>Electronic structure (drawn)</th>
</tr>
</thead>
<tbody>
<tr>
<td>He</td>
<td>2</td>
<td>2</td>
<td>2</td>
<td><img src="image" alt="He structure" /></td>
</tr>
<tr>
<td>F</td>
<td>9</td>
<td>9</td>
<td>2,7</td>
<td><img src="image" alt="F structure" /></td>
</tr>
<tr>
<td>Al</td>
<td>13</td>
<td>13</td>
<td>2,8,3</td>
<td><img src="image" alt="Al structure" /></td>
</tr>
<tr>
<td>K</td>
<td>19</td>
<td>19</td>
<td>2,8,8,1</td>
<td><img src="image" alt="K structure" /></td>
</tr>
</tbody>
</table>

**Ions**

**Ions** are particles with an electric charge because they do **not** contain the same number of protons and electrons. Remember that protons are positive and electrons are negative. Positive ions have more protons than electrons. Negative ions have more electrons than protons.

For example:
- An ion with 11 protons (total charge 11+) and 10 electrons (total charge 10−) will have an overall charge of 1+.
- An ion with 16 protons (total charge 16+) and 18 electrons (total charge 18−) will have an overall charge of 2−.

**Test yourself**

16 Write the electronic structure of the following atoms: $\text{O}$, $\text{Na}$, $\text{Ca}$.
17 Lithium atoms contain 3 electrons and have the electronic structure 2,1. State why the electrons are not all in the first shell.

**Show you can...**

The diagram shows an atom of an element X, where: $\text{e}$ represents an electron; $\text{n}$ represents a neutron; and $\text{p}$ represents a proton.

a) Name the element X.
b) Write the electronic structure of X.
c) What is the mass number of this atom of element X?
d) Name the part of the atom shaded red.
Table 9.5 shows some common ions.

<table>
<thead>
<tr>
<th>Ion</th>
<th>Li⁺</th>
<th>Al³⁺</th>
<th>Cl⁻</th>
<th>O²⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atomic number</td>
<td>3</td>
<td>13</td>
<td>17</td>
<td>8</td>
</tr>
<tr>
<td>Number of protons</td>
<td>3 (charge 3+)</td>
<td>13 (charge 13+)</td>
<td>17 (charge 17+)</td>
<td>8 (charge 8+)</td>
</tr>
<tr>
<td>Number of electrons</td>
<td>2 (charge 2−)</td>
<td>10 (charge 10−)</td>
<td>18 (charge 18−)</td>
<td>10 (charge 10−)</td>
</tr>
<tr>
<td>Overall charge</td>
<td>1+</td>
<td>3+</td>
<td>1−</td>
<td>2−</td>
</tr>
<tr>
<td>Electronic structure (written)</td>
<td>2</td>
<td>2,8</td>
<td>2,8,8</td>
<td>2,8</td>
</tr>
<tr>
<td>Electronic structure (drawn)</td>
<td>![Diagram of Li⁺]</td>
<td>![Diagram of Al³⁺]</td>
<td>![Diagram of Cl⁻]</td>
<td>![Diagram of O²⁻]</td>
</tr>
</tbody>
</table>

Simple ions (those made from single atoms) have the same electronic structure as the elements in Group 0 of the periodic table (Table 9.6). The elements in Group 0 are called the noble gases. The noble gases have very stable electronic structures.

Table 9.6 Common ions with the same electronic structure as Group 0 elements.

<table>
<thead>
<tr>
<th>Group 0 element</th>
<th>He</th>
<th>Ne</th>
<th>Ar</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electronic structure</td>
<td>2</td>
<td>2,8</td>
<td>2,8,8</td>
</tr>
<tr>
<td>Common ions with the same electronic structure</td>
<td>Li⁺, Be²⁺</td>
<td>O²⁻, F⁻, Na⁺, Mg²⁺, Al³⁺</td>
<td>S²⁻, Cl⁻, K⁺, Ca²⁺</td>
</tr>
</tbody>
</table>

The hydrogen ion (H⁺) is the only simple ion that does not have the electronic structure of a noble gas. It does not have any electrons at all. This makes it a very special ion with special properties, and it is the H⁺ ion that is responsible for the behaviour of acids.

Show you can...

Test yourself

18 What is the charge of a particle with 19 protons and 18 electrons?
19 What is the charge of a particle with 7 protons and 10 electrons?
20 What is the electronic structure of the P³⁻ ion?
21 How many protons, neutrons and electrons are there in the ¹⁷F⁻ ion?
22 What is the link between the electronic structure of ions and the Group 0 elements (the noble gases)?
Development of ideas about the structure of atoms

The idea that everything was made of particles called atoms was accepted in the early 1800s after work by John Dalton. At that time, however, people thought that atoms were the smallest possible particles and the word atom comes from the Greek word atomos which means something that cannot be divided.

However, in 1897 the electron was discovered by J.J. Thompson while carrying out experiments on the conduction of electricity through gases. He discovered that electrons were tiny, negatively charged particles that were much smaller and lighter than atoms. He came up with what was called the ‘plum pudding’ model of the atom. In this model, the atom was a ball of positive charge with the negative electrons spread through the atom (Figure 9.5).

A few years later in 1911, this model was replaced following some remarkable work from Hans Geiger and Ernest Marsden working with Ernest Rutherford. They fired alpha particles (He\(^{2+}\) ions) at a very thin piece of gold foil. They expected the particles to pass straight through the foil but a tiny fraction were deflected or even bounced back. This did not fit in with the plum pudding model. Rutherford worked out that the scattering of some of the alpha particles meant that there must be a tiny, positive nucleus at the centre of each atom. This new model was known as the nuclear model (Figure 9.6).

In 1913, Neils Bohr adapted the nuclear model to suggest that the electrons moved in stable orbits at specific distances from the nucleus called shells. Bohr’s theoretical calculations agreed with observations from experiments.

Further experiments led to the idea that the positive charge of the nucleus was made up from particles which were given the name protons.
Scientists realised that there was some mass in atoms that could not be explained by this model, and in 1932 James Chadwick discovered a new particle inside the nucleus that had the same mass as a proton but had no electric charge. This particle was given the name neutron.

The model has been developed further since then, but the basic idea of atoms being made up of a tiny central nucleus containing protons and neutrons surrounded by electrons in shells remains (Figure 9.7).

The development of ideas about atomic structure shows very well how scientific models and theories develop over time. When new discoveries are made, models and theories may have to be altered or sometimes completely replaced if they do not fit in with the new discoveries.

### Test yourself

23 What was discovered that led to scientists realising that atoms were made up of smaller particles?

24 Why was the plum pudding model replaced?

25 Why would a nucleus deflect an alpha particle?

---

### Reactions of elements

#### Elements in the periodic table

An element is a substance containing only one type of atom. For example, in the element carbon all the atoms are carbon atoms meaning that all the atoms have 6 protons and so have the atomic number 6. Elements cannot be broken down into simpler substances.

Atoms are known with atomic numbers ranging from 1 to just over 100. All the elements are listed in the periodic table. The elements are listed in order of atomic number (Figure 9.8).

Atoms of each element are given their own symbol, each with one, two or three letters. The first letter is always a capital letter with any further letters being small letters. For example, carbon has the symbol C while copper has the symbol Cu.
### TIPS

- Metals are found to the left and towards the bottom of the periodic table.
- Non-metals are found towards the right and top of the periodic table.

### Metals and non-metals

Over three-quarters of the elements are metals, with most of the rest being non-metals. Typical properties of metals and non-metals are shown in Table 9.8, although there are some exceptions.

<table>
<thead>
<tr>
<th>Group</th>
<th>Metals</th>
<th>Non-metals</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Li, Be, Na</td>
<td>C, O, N</td>
</tr>
<tr>
<td>2</td>
<td>K, Ca, Sc</td>
<td>S, Cl</td>
</tr>
<tr>
<td>3</td>
<td>Rb, Sr, Y</td>
<td>Ar, Kr</td>
</tr>
<tr>
<td>4</td>
<td>Cs, Ba, La</td>
<td>Xe, Rn</td>
</tr>
<tr>
<td>5</td>
<td>Fr, Ra, Ac</td>
<td></td>
</tr>
<tr>
<td>6</td>
<td></td>
<td></td>
</tr>
<tr>
<td>7</td>
<td></td>
<td></td>
</tr>
<tr>
<td>8</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Table 9.8 Properties of metals and non-metals.

<table>
<thead>
<tr>
<th>Metals</th>
<th>Non-metals</th>
</tr>
</thead>
<tbody>
<tr>
<td>Melting and boiling points</td>
<td>High</td>
</tr>
<tr>
<td>Conductivity</td>
<td>Thermal and electrical conductor</td>
</tr>
<tr>
<td>Density</td>
<td>High density</td>
</tr>
<tr>
<td>Appearance</td>
<td>Shiny when polished</td>
</tr>
<tr>
<td>Malleability</td>
<td>Can be hammered into shape</td>
</tr>
<tr>
<td>Reaction with non-metals</td>
<td>React to form positive ions in ionic compounds</td>
</tr>
<tr>
<td>Reaction with metals</td>
<td>No reaction</td>
</tr>
<tr>
<td>Acid-base properties of oxides</td>
<td>Metal oxides are basic</td>
</tr>
</tbody>
</table>

There are a few elements around the dividing line between metals and non-metals, such as silicon and germanium, that are hard to classify as they have some properties of metals and some of non-metals.
### Test yourself

26 Is each of the following elements a metal or non-metal?

- a) Element 1 is a dull solid at room temperature that easily melts when warmed.
- b) Element 2 is a dense solid that is a thermal conductor.
- c) Element 3 reacts with oxygen to form an oxide which dissolves in rain water to form acid rain.
- d) Element 4 reacts with chlorine to form a compound made of molecules.
- e) Element 5 reacts with sodium to form a compound made of ions.

### Show you can...

Figure 9.9 shows magnesium and oxygen reacting to form a single product.

- a) State two differences in the physical properties of magnesium and oxygen.
- b) Suggest the name of the product of this reaction.
- c) Is the product acidic or basic?
- d) Does the product consist of ions or molecules?

### Reactions between elements

When elements react with each other they form compounds. **Compounds** are substances made from different elements bonded (chemically joined) together. A chemical reaction takes place when elements combine to form compounds. Chemical reactions always involve the formation of one or more new substances and there is usually a detectable energy change.

When elements react with each other, electrons are either shared with other elements or transferred from one element to another. This is done so that atoms obtain the stable electronic structure of the noble gases (Group 0 elements). Bonding will be covered in more detail in the next chapter.

Table 9.9 shows what happens in general when elements react with each other.

<table>
<thead>
<tr>
<th>Elements reacting</th>
<th>What happens to the electrons to obtain noble gas electronic structures</th>
<th>Type of particles formed</th>
<th>Type of compound formed</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Non-metal + non-metal</td>
<td>Electrons shared</td>
<td>Molecules (where atoms are joined to each other by covalent bonds)</td>
<td>Molecular compound</td>
<td>Hydrogen reacts with oxygen by sharing electrons and forming molecules of water</td>
</tr>
<tr>
<td>Metal + non-metal</td>
<td>Electrons transferred from metal to non-metal</td>
<td>Positive and negative ions</td>
<td>Ionic compound</td>
<td>Sodium reacts with chlorine by transferring electrons from sodium to chlorine to form sodium chloride which is made of ions</td>
</tr>
<tr>
<td>Metal + metal</td>
<td>No reaction as both metals cannot lose electrons</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

### KEY TERM

**Compound** Substance made from different elements chemically bonded together.
Test yourself

27 Do the following elements react with each other, and if they do, what type of compound is formed?

a) potassium + oxygen
b) bromine + iodine
c) oxygen + sulfur
d) sulfur + magnesium
e) calcium + potassium
f) nitrogen + hydrogen

Show you can...

The electronic structures of the atoms of 5 different elements, A, B, C, D and E, are shown below.

A 2,8,8     B 2,8,8,1     C 2,6
D 2,1       E 2,8,7

Using the letters A, B, C, D or E choose:

a) An unreactive element.
b) Two elements found in the same Group of the periodic table.
c) An element whose atoms will form ions with a charge of 2−.
d) Two elements that react to form an ionic compound.
e) Two elements that react to form a covalent compound.

The periodic table

Electronic structure and the periodic table

The elements are placed in the periodic table in order of increasing atomic number (the number of protons). Figure 9.10 shows the first 36 elements in the periodic table.

Figure 9.10 The first 36 elements in the periodic table, with their atomic (proton) number.
The table can be seen as arranging the elements by electronic structure. At the end of each period, a noble gas stable electronic structure is reached and then a new energy level (shell) starts to be filled at the start of the next period. The electronic structure of the first 20 elements is shown in Figure 9.11. Note the increasing number of electrons in the elements going from left to right in each row (period).

Elements in the same group (column) have the same number of electrons in their outer shell. The number of electrons in the outer shell equals the Group number. For example, all the elements in Group 1 have 1 electron in their outer shell (Li = 2,1; Na = 2,8,1; K = 2,8,8,1) (Figure 9.11). All the elements in Group 7 have 7 electrons in their outer shell (F = 2,7; Cl = 2,8,7). The only exception to this is Group 0 where all the elements have 8 electrons in their outer shell except helium which has 2 electrons (but the first shell can only hold 2 electrons).

All the elements in the same group have similar chemical properties because they have the same number of electrons in their outer shell.

Elements in the same period (row) have the same number of shells (electron levels) – see Figure 9.11.

The main elements of Group 0 are helium, neon, argon, krypton, xenon and radon (Figure 9.12 and Table 9.10). They are known as the noble gases (Table 9.11). These atoms all have stable electronic structures. Helium’s outer shell is full with 2 electrons while the others have 8 electrons in their outer shells.
Figure 9.12 Group 0 – the noble gases.

Table 9.10 The noble gases.

<table>
<thead>
<tr>
<th>Element</th>
<th>Formula</th>
<th>Appearance at room temperature</th>
<th>Number of electrons in outer shell</th>
<th>Relative mass of atoms</th>
<th>Boiling point in °C</th>
</tr>
</thead>
<tbody>
<tr>
<td>Helium</td>
<td>He</td>
<td>Colourless gas</td>
<td>2</td>
<td>4</td>
<td>−269</td>
</tr>
<tr>
<td>Neon</td>
<td>Ne</td>
<td>Colourless gas</td>
<td>8</td>
<td>20</td>
<td>−246</td>
</tr>
<tr>
<td>Argon</td>
<td>Ar</td>
<td>Colourless gas</td>
<td>8</td>
<td>40</td>
<td>−190</td>
</tr>
<tr>
<td>Krypton</td>
<td>Kr</td>
<td>Colourless gas</td>
<td>8</td>
<td>84</td>
<td>−157</td>
</tr>
<tr>
<td>Xenon</td>
<td>Xe</td>
<td>Colourless gas</td>
<td>8</td>
<td>131</td>
<td>−111</td>
</tr>
<tr>
<td>Radon</td>
<td>Rn</td>
<td>Colourless gas</td>
<td>8</td>
<td>222</td>
<td>−62</td>
</tr>
</tbody>
</table>

Table 9.11 Properties of the noble gases.

- **Metals or non-metals?** All the elements are non-metals.
- **Boiling points** The noble gases are all colourless gases with low boiling points. The boiling points increase as the atoms get heavier going down the group.
- **Reactivity** The Group 0 elements are very unreactive and do not easily react to form molecules or ions because their atoms have stable electron arrangements.

Show you can…

Copy and complete the table:

Table 9.12

<table>
<thead>
<tr>
<th>Element</th>
<th>Reactive or unreactive?</th>
<th>Metal or non-metal?</th>
<th>Solid, liquid or gas at room temperature?</th>
<th>Electron structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>He</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ar</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Test yourself**

31 **Why are the noble gases unreactive?**
32 **Suggest a reason why the noble gases are referred to as being in Group 0 rather than Group 8.**
33 **Some atoms of element 118 (Uuo) have been produced. Element 118 is in Group 0. Predict the chemical and physical properties of this element.**

**Group 1 – the alkali metals**

The main elements of Group 1 are lithium, sodium, potassium, rubidium and caesium (Table 9.13). They are known as the alkali metals (Figure 9.13 and Table 9.14). The Group 1 elements have similar chemical and physical properties because they all have one electron in their outer shell. They are all soft metals that can be cut with a knife (Figure 9.14). They are very reactive and so are stored in bottles of oil to stop them reacting with water and oxygen. They are very reactive because they have only one electron in their outer shell, which can easily be lost.
The alkali metals are all soft and can be cut with a knife.

Table 9.13 The alkali metals.

<table>
<thead>
<tr>
<th>Element</th>
<th>Formula</th>
<th>Appearance at room temperature</th>
<th>Number of electrons in outer shell</th>
<th>Relative mass of atoms</th>
<th>Melting point in °C</th>
<th>Density in g/cm³</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lithium</td>
<td>Li</td>
<td>Silvery-grey metal</td>
<td>1</td>
<td>7</td>
<td>180</td>
<td>0.53</td>
</tr>
<tr>
<td>Sodium</td>
<td>Na</td>
<td>Silvery-grey metal</td>
<td>1</td>
<td>23</td>
<td>98</td>
<td>0.97</td>
</tr>
<tr>
<td>Potassium</td>
<td>K</td>
<td>Silvery-grey metal</td>
<td>1</td>
<td>39</td>
<td>63</td>
<td>0.89</td>
</tr>
<tr>
<td>Rubidium</td>
<td>Rb</td>
<td>Silvery-grey metal</td>
<td>1</td>
<td>85</td>
<td>39</td>
<td>1.53</td>
</tr>
<tr>
<td>Caesium</td>
<td>Cs</td>
<td>Silvery-grey metal</td>
<td>1</td>
<td>133</td>
<td>28</td>
<td>1.93</td>
</tr>
</tbody>
</table>

All the elements are metals.

The alkali metals are all solids with relatively low melting points at room temperature. The melting points decrease as the atoms get bigger going down the group.

Lithium, sodium and potassium all float on water as they are less dense than water.

The metals all react easily with non-metals by the transfer of electrons from the metal to the non-metal forming compounds made of ions. Alkali metals always form 1+ ions (e.g. Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺) as they have one electron in their outer shell which they lose when they react to obtain a noble gas electronic structure.

The metals all burn in oxygen to form metal oxides which are white powders:
- metal + oxygen → metal oxide
- e.g. Na + O₂ → Na₂O

The metals burn with different colour flames. For example lithium burns with a crimson-red flame, sodium with a yellow-orange flame and potassium with a lilac flame.

The alkali metals all burn in chlorine to form metal chlorides which are white powders:
- metal + chlorine → metal chloride
- e.g. sodium + chlorine → sodium chloride
- e.g. 2Na + Cl₂ → 2NaCl

The alkali metals all react with water, releasing hydrogen gas and forming a solution containing a metal hydroxide:
- metal + water → metal hydroxide + hydrogen
- e.g. sodium + water → sodium hydroxide + hydrogen
- e.g. 2Na + 2H₂O → 2NaOH + H₂

The solution of the metal hydroxide that is formed is alkaline.

Compounds made from Group 1 metals

Compounds made from alkali metals:
- Ionic
- White solids
- Dissolve in water to form colourless solutions
Reactivity trend of the alkali metals

The alkali metals get more reactive the further down the group (Figure 9.15). This can be seen when the alkali metals react with water.

<table>
<thead>
<tr>
<th>Description</th>
<th>Lithium</th>
<th>Potassium</th>
<th>Caesium</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fizzes, moves on the surface of the water</td>
<td>Fizzes, melts, burns with lilac flame, moves on the surface of the water</td>
<td>Explosive reaction</td>
<td></td>
</tr>
</tbody>
</table>

When the alkali metals react they are losing their outer shell electron in order to get a stable electronic structure. The further down the group, the further away the outer electron is from the nucleus as the atoms get bigger. This means that the outer electron is less strongly attracted to the nucleus and so easier to lose. The easier the electron is to lose, the more reactive the alkali metal.

Test yourself

34 Why are the alkali metals reactive?
35 Write a word equation for the reaction of potassium with water.
36 Explain why the solution formed when potassium reacts with water has a high pH.
37 Potassium reacts with chlorine to form an ionic compound. Explain why this reaction happens.
38 Explain why potassium is more reactive than sodium.
39 Francium is the last element in Group 1. Predict the chemical and physical properties of francium.

Group 7 – the halogens

The main elements of Group 7 are fluorine, chlorine, bromine and iodine (Figure 9.16 and Table 9.15). They are known as the halogens (Figure 9.16 and Table 9.16). The Group 7 elements have similar chemical and physical properties because they all have seven electrons in their outer shell. The halogens are reactive because they only need to gain one electron to gain a noble gas electronic structure. The particles in each of the elements are molecules containing two atoms (diatomic molecules), such as F₂, Cl₂, Br₂ and I₂.
Table 9.15 The halogens.

<table>
<thead>
<tr>
<th>Element</th>
<th>Formula</th>
<th>Appearance at room temperature</th>
<th>Number of electrons in outer shell</th>
<th>Relative mass of molecules</th>
<th>Melting point in °C</th>
<th>Boiling point in °C</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fluorine</td>
<td>F₂</td>
<td>Pale yellow gas</td>
<td>7</td>
<td>38</td>
<td>−220</td>
<td>−188</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl₂</td>
<td>Pale green gas</td>
<td>7</td>
<td>71</td>
<td>−102</td>
<td>−34</td>
</tr>
<tr>
<td>Bromine</td>
<td>Br₂</td>
<td>Dark brown liquid</td>
<td>7</td>
<td>160</td>
<td>−7</td>
<td>59</td>
</tr>
<tr>
<td>Iodine</td>
<td>I₂</td>
<td>Grey solid</td>
<td>7</td>
<td>254</td>
<td>114</td>
<td>184</td>
</tr>
</tbody>
</table>

Table 9.16 Properties of the halogens.

<table>
<thead>
<tr>
<th>Metals or non-metals?</th>
<th>Fluorine, chlorine, bromine and iodine are all <strong>non-metals</strong>.</th>
</tr>
</thead>
<tbody>
<tr>
<td>Toxicity</td>
<td>Each of the halogens is toxic.</td>
</tr>
<tr>
<td>Melting and boiling points</td>
<td>The halogens have low melting and boiling points. The melting and boiling points increase as the molecules get heavier going down the group.</td>
</tr>
<tr>
<td>Reaction with non-metals</td>
<td>The halogens react with other non-metals by sharing electrons to form compounds made of molecules (molecular compounds).</td>
</tr>
<tr>
<td>Reaction with metals</td>
<td>The halogens all react easily with metals by the transfer of electrons from the metal to the halogen forming compounds made of ions (ionic compounds). Halogens always form <strong>1− ions</strong> (e.g. F−, Cl−, Br−, I−, all known as <strong>halide ions</strong>) as they have seven electrons in their outer shell and gain one more electron when they react to get a noble gas electronic structure.</td>
</tr>
</tbody>
</table>

**Reactivity trend of the halogens**

The halogens get less reactive the further down the group. This can be seen by looking at which halogens can displace each other from compounds. Compounds containing halogens, such as sodium chloride and potassium bromide are often called **halides** or **halide compounds**.

A more reactive element can displace a less reactive element from a compound. You have seen this (before GCSE) with metals when a more reactive metal can displace a less reactive metal from a compound. For example, aluminium can displace iron from iron oxide because aluminium is more reactive than iron.

\[
aluminium + \text{iron oxide} \rightarrow \text{aluminium oxide} + \text{iron}
\]

In a similar way, a more reactive non-metal can displace a less reactive non-metal from a compound. This means that a more reactive halogen can displace a less reactive halogen from a halide compound.
This can be seen when aqueous solutions of the halogens react with aqueous solutions of halide compounds (aqueous means dissolved in water) (Table 9.17 and Figure 9.18).

Table 9.17 Displacement reactions involving halogens and halide compounds.

<table>
<thead>
<tr>
<th>Potassium chloride (aq)</th>
<th>Bromine (aq)</th>
<th>Iodine (aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Chlorine (aq)</td>
<td>No reaction</td>
<td>No reaction</td>
</tr>
<tr>
<td>Potassium bromide (aq)</td>
<td>chlorine + potassium bromide $\rightarrow$ potassium chloride + bromine Cl$_2$ + 2KBr $\rightarrow$ 2KCl + Br$_2$ (Cl$_2$ + 2Br$^-$ $\rightarrow$ 2Cl$^-$ + Br$_2$) Yellow solution formed (due to production of bromine) Chlorine displaces bromine</td>
<td>No reaction</td>
</tr>
<tr>
<td>Potassium iodide (aq)</td>
<td>chlorine + potassium iodide $\rightarrow$ potassium chloride + iodine Cl$_2$ + 2KI $\rightarrow$ 2KCl + I$_2$ (Cl$_2$ + 2I$^-$ $\rightarrow$ 2Cl$^-$ + I$_2$) Brown solution formed (due to production of iodine) Chlorine displaces iodine</td>
<td>bromine + potassium iodide $\rightarrow$ potassium bromide + iodine Br$_2$ + 2KI $\rightarrow$ 2KBr + I$_2$ (Br$_2$ + 2I$^-$ $\rightarrow$ 2Br$^-$ + I$_2$) Brown solution formed (due to production of iodine) Bromine displaces iodine</td>
</tr>
</tbody>
</table>

It can be seen from these reactions that the trend in reactivity for these three halogens is:

Most reactive Chlorine

Least reactive Iodine

In general, the **further down the group the less reactive the halogen** (Figure 9.19). The higher up the group, the more reactive the halogen. This means that fluorine is the most reactive halogen. You will not do experiments with fluorine because it is very reactive and toxic.

When the halogens react they gain one electron in order to get a noble gas electronic structure. The further down the group, the electron gained will enter an energy level further away from the nucleus as the atoms get bigger. This means that the electron gained is less strongly attracted to the nucleus and so harder to gain. The harder the electron is to gain, the less reactive the halogen.

**Test yourself**

40 Why are the halogens reactive?
41 All the halogens are made of diatomic molecules. What are diatomic molecules?
42 Predict what would happen, and why, if fluorine and sodium chloride were mixed.
43 Bromine reacts with chlorine to form a molecular compound. Explain why this reaction happens.
44 Explain why chlorine is more reactive than bromine.

**TIP**

The explanation for the reactivity trend in Group 7 is about the distance between the nucleus and the electron gained which is from **outside** the atom – it is not about the outer shell electrons.
History of the periodic table

As more elements were discovered, scientists tried to classify the elements into some sort of order and pattern. This was originally done before the discovery of protons, neutrons and electrons. Scientists’ first attempts were based on the use of the atomic weights of elements which we now know as relative atomic mass.

John Newlands spotted that the properties of elements seemed to repeat every eighth element when placed in order of atomic weight. He called this the ‘law of octaves’ as it was similar to notes in musical scales. One of the successes of his table was that he had lithium, sodium and potassium in the same group, each of which has very similar properties (Figure 9.22).
At the time though, only about 50 elements were known and his table was not accepted because it only worked for the first 20 or so of the known elements. After that there were problems, such as copper being in the same group as lithium, sodium and potassium. Copper has very different properties to those metals. For example, copper does not react with water but the other three react vigorously with water.

A few years later in 1869, a Russian chemist called Dmitri Mendeleev devised a table which has become the basis for the periodic table today (Figure 9.23). He also placed elements in order of atomic weight, but crucially he did two things differently to Newlands.

1 Mendeleev left gaps for elements he predicted had yet to be discovered. He also predicted the properties of these elements.

2 Mendeleev was prepared to alter slightly the order of the elements if he thought it fitted the properties better. For example, he swapped around the order of iodine (atomic weight 127) and tellurium (atomic weight 128). He placed iodine after tellurium so that it was in the same group as fluorine, chlorine and bromine which have very similar properties. Mendeleev actually believed that the atomic weights must have been measured incorrectly.
Over the next few years, elements were discovered that Mendeleev had predicted would exist (Table 9.19). These included gallium (1875), scandium (1879) and germanium (1886). In each case the properties of the element closely matched Mendeleev’s predictions. Table 9.19 shows some of the properties that Mendeleev predicted for the element he called eka-silicon and that we call germanium.

### Table 9.19 Mendeleev’s predictions for eka-silicon (germanium).

<table>
<thead>
<tr>
<th></th>
<th>Element</th>
<th>Appearance</th>
<th>Atomic weight</th>
<th>Melting point in °C</th>
<th>Density in g/cm³</th>
<th>Formula of oxide</th>
<th>Formula of chloride</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mendeleev’s predictions</td>
<td>Eka-silicon (Es)</td>
<td>Grey metal</td>
<td>72</td>
<td>high</td>
<td>5.5</td>
<td>EsO₂</td>
<td>EsCl₄</td>
</tr>
<tr>
<td>Actual properties</td>
<td>Germanium (Ge)</td>
<td>Grey-white metal</td>
<td>73</td>
<td>947</td>
<td>5.4</td>
<td>GeO₂</td>
<td>GeCl₄</td>
</tr>
</tbody>
</table>

As these elements were discovered, Mendeleev’s ideas and table became well accepted and formed the basis of the periodic table as we know it today. We now know that Mendeleev placed the elements in order of atomic number (the number of protons in an atom) even though he did not know about the existence of protons. It is the atomic number rather than atomic weight that matters, because elements are made of a mixture of isotopes and, depending on the isotope, the atomic weight will differ.

The story of Mendeleev illustrates how strong support can come for a scientific idea if predictions made using that theory are later found to be correct.

### Mixtures

A mixture consists of two or more substances that are mixed together and not chemically combined. In a mixture, each substance has its own properties. Mixtures are very different from compounds (Table 9.20).

### Table 9.20 Differences between compounds and mixtures.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Description</th>
<th>Proportions</th>
<th>Separation</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>A substance made from two or more elements chemically bonded together. A compound is a single substance with its own unique properties</td>
<td>Each compound has a fixed proportion of elements (so each compound has a fixed formula)</td>
<td>Compounds can only be separated back into elements by chemical reaction because the elements are chemically joined.</td>
</tr>
<tr>
<td></td>
<td>Two or more substances each with their own properties (the different substances are not chemically joined to each other)</td>
<td>There can be any amount of each substance in the mixture</td>
<td>No chemical reaction is needed as the substances in the mixture are not chemically joined. They can be separated by physical methods (e.g. filtration, distillation).</td>
</tr>
</tbody>
</table>
Sodium is a very reactive, dangerous, grey metal that reacts vigorously with water. Chlorine is a pale green, toxic gas that is very reactive. In a mixture of sodium and chlorine each substance keeps its own properties as a grey metal and green gas, respectively. It is easy to separate the sodium and chlorine because they are not chemically joined together.

However, if heated together sodium reacts with chlorine to make the compound sodium chloride. Sodium chloride is very different from both sodium and chlorine. Sodium chloride is a white solid that is not very reactive and is safe to eat. It is very difficult to break sodium chloride back down into the elements because the sodium and chlorine are chemically joined together.

<table>
<thead>
<tr>
<th>Table 9.21</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium</td>
</tr>
<tr>
<td>Grey, highly reactive, dangerous metal</td>
</tr>
</tbody>
</table>

**Show you can...**

For each of the substances A, B, C, D decide if it is an element, compound or mixture. If any substance is a mixture decide if it is a mixture of elements, a mixture of elements and compounds, or a mixture of compounds.

---

**Separating mixtures**

The substances in a mixture are quite easy to separate because the substances are not chemically joined to each other. Different methods are used depending on what type of mixture there is (Table 9.22).

<table>
<thead>
<tr>
<th>Table 9.22 Different methods of separating mixtures.</th>
</tr>
</thead>
<tbody>
<tr>
<td>Type of mixture</td>
</tr>
<tr>
<td>Method of separation</td>
</tr>
</tbody>
</table>

---

**TIP**

Some definitions of key words:

- **solute**: the solid substance that dissolves in a solvent
- **solvent**: the liquid that a solute dissolves in
- **solution**: a solute dissolved in a solvent
- **soluble**: when a substance will dissolve in a solvent
- **insoluble**: when a substance does not dissolve in a solvent.
Filtration
This method is used to separate an insoluble solid from a liquid. For example, it could be used to separate sand from water.

The mixture is poured through a funnel containing a piece of filter paper. The liquid (called the filtrate) passes through the paper and the solid (called the residue) remains on the filter paper (Figure 9.25).

Evaporation
This method is used to separate a dissolved solid from the solvent it is dissolved in. For example, it could be used to separate salt from water.

The mixture is placed in an evaporating dish and heated until all the solvent has evaporated or boiled, leaving the solid in the evaporating basin (Figure 9.26).

Crystallisation
This method is also used to separate a dissolved solid from the solvent it is dissolved in. For example, it could be used to separate copper sulfate crystals from a solution of copper sulfate (Figure 9.27).

The mixture is heated to boil off some of the solvent to create a hot, saturated solution. A saturated solution is one in which no more solute can dissolve at that temperature. As it cools down, the solute becomes less soluble and so cannot remain dissolved, so some of the solute crystallises out of the solution as crystals. The crystals can then be separated from the rest of the solution by filtration.

Simple distillation
This method is used to separate the solvent from a solution. For example, it could be used to separate pure water from sea water.

The mixture is heated and the solvent boils. The vaporised solvent passes through a water-cooled condenser where it cools and condenses. The condenser directs the condensed solvent into a container away from the original solution (Figure 9.28).
Mixtures

Fractional distillation

Liquids that mix together are called **miscible** liquids. Water and alcohol are examples of miscible liquids. Fractional distillation is used to separate mixtures of miscible liquids. It works because the liquids have different boiling points.

The apparatus used is similar to that for simple distillation, but a long column (called a fractionating column) is used to help separate different liquids as they boil. The fractionating column often contains glass beads.

In industry, such as in the fractional distillation of crude oil, the whole mixture is vaporised and then condensed in a fractionating column which is hot at the bottom and cold at the top. The liquids will condense at different heights in the fractionating column (Figure 9.29).

Separating funnel

Liquids that do not mix together are called **immiscible** liquids. Oil and water is an example of liquids that are immiscible with each other. They can be separated in a **separating funnel**. The liquids form two layers and the bottom layer can be removed using the tap at the bottom of the funnel. The liquid with the greater density is the lower layer (Figure 9.30).

Chromatography

There are many forms of chromatography. Paper chromatography is used to separate mixtures of substances dissolved in a solvent.

A piece of chromatography paper, with the mixture on, is placed upright in a beaker so that the bottom of the paper is in the solvent. Over time, the solvent soaks up the paper. The substances move up the paper at different speeds and so are separated (Figure 9.31).

**Test yourself**

How would you separate the following mixtures?

a) alcohol from a mixture of alcohol and water
b) magnesium hydroxide from a suspension of insoluble magnesium hydroxide in water
c) pure dry cleaning solvent from waste dry cleaning solvent containing dirt that dissolved in the solvent from clothes
d) sunflower oil and water
e) food colourings in a sweet.

**KEY TERMS**

**Miscible** Liquids that mix together.

**Immiscible** Liquids that do not mix together and separate into layers.

**Separating funnel** Glass container with a tap used to separate immiscible liquids.
Filtration, Distillation, Fractional Distillation

For each of these separation methods pick two words or phrases from the list and insert them into a copy of the table with an explanation of their meaning. Also include the type of mixture separated by each method:

condenser, distillate, fractionating column, filtrate, miscible liquids, residue.

Table 9.23

<table>
<thead>
<tr>
<th>Type of mixture separated</th>
<th>Filtration</th>
<th>Distillation</th>
<th>Fractional distillation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Important word and definition</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Important word and definition</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Show you can...**

Three common methods of separation include filtration, distillation and fractional distillation.

For each of these separation methods pick two words or phrases from the list and insert them into a copy of the table with an explanation of their meaning. Also include the type of mixture separated by each method:

condenser, distillate, fractionating column, filtrate, miscible liquids, residue.

**Rock salt**

Common salt is sodium chloride and is found naturally in large amounts in seawater or in underground deposits. Sodium chloride can be extracted from underground by the process of solution mining.

**Method:**

i. Place 8 spatulas of rock salt into a mortar and grind using a pestle.

ii. Place the rock salt into a beaker and quarter fill with water.

iii. Place on a gauze and tripod and heat, stirring with a glass rod. Stop heating when the salt has dissolved – the sand and clay will be left undissolved.

iv. Allow to cool and then filter.

v. Heat until half the volume of liquid is left.

vi. Place the evaporating basin on the windowsill to evaporate off the rest of the water slowly. Pure salt crystals should be left.

**Figure 9.32**

1 a) On what physical property of sodium chloride does this process depend?

b) Suggest one reason why solution mining uses a lot of energy.

c) Suggest one negative effect which solution mining has on the environment.

d) Suggest how sodium chloride is obtained from the concentrated salt solution.

2 Rock salt is a mixture of salt, sand and clay. To separate pure salt from rock salt, the method listed below can be used in the laboratory.

**Method:**

i. Place 8 spatulas of rock salt into a mortar and grind using a pestle.

ii. Place the rock salt into a beaker and quarter fill with water.

iii. Place on a gauze and tripod and heat, stirring with a glass rod. Stop heating when the salt has dissolved – the sand and clay will be left undissolved.

iv. Allow to cool and then filter.

v. Heat until half the volume of liquid is left.

vi. Place the evaporating basin on the windowsill to evaporate off the rest of the water slowly. Pure salt crystals should be left.

Choose one step of the method (i to vi) which is best represented in each photograph A–C.

3 a) Why is rock salt considered to be a mixture?

b) What was the purpose of grinding the rock salt?

c) Why was the mixture heated and stirred?

d) State what the filtrate contains.

e) State what the residue contains.

f) Explain why the salt obtained may still be contaminated with sand and suggest how you would improve your experiment to obtain a purer sample of salt.
Chapter review questions

1 Choose from the following list of elements to answer the questions below:
   bromine  calcium  krypton  nickel  nitrogen  potassium  silicon
   a) Which element is most like lithium?
   b) Which element is most like iron?
   c) Which element is most like helium?
   d) Which element is most like fluorine?
   e) Which element is most like carbon?

2 In which group or area of the periodic table would you find these elements?
   a) Element A has 7 electrons in its outer shell.
   b) Element B reacts vigorously with water to give off hydrogen gas and an alkaline solution.
   c) Element C is a metal with 4 electrons in its outer shell.
   d) Element D is a colourless gas that does not react at all.
   e) Element E forms coloured compounds.
   f) Element F is toxic and is made of diatomic molecules.
   g) Element G forms 1− ions when it reacts with metals to form ionic compounds.
   h) Element H can form both 1+ and 2+ ions.
   i) Element I is a metal that floats on water.
   j) Element J has the electronic structure 2,8,8,6.
   k) Element K has 12 protons.
   l) Element L has a full outer shell.
   m) Element M can act as a catalyst.

3 Identify a mixture that could be separated by each of the following methods.
   a) simple distillation
   b) filtration
   c) crystallisation
   d) evaporation
   e) chromatography
   f) fractional distillation

4 Look at the following atoms and ions.
   \[ {}^{12}\text{C} \quad {}^{14}\text{C} \quad {}^{16}\text{O}^{2−} \quad {}^{19}\text{F}− \quad {}^{20}\text{Ne} \]
   Which of these atoms and ions, if any,
   a) are isotopes?
   b) have 9 protons?
   c) have 10 electrons?
   d) have 10 neutrons?
   e) have more protons than electrons?

5 Caesium atoms are among the largest atoms. A caesium atom has a radius of 0.260nm. Write this in metres in standard form.
6 Predict whether each of the following pairs of elements will (i) react by sharing electrons; (ii) react by transferring electrons; or (iii) not react:

   a) magnesium + oxygen
   b) sulfur + hydrogen
   c) aluminium + magnesium
   d) argon + oxygen
   e) bromine + phosphorus
   f) fluorine + lithium

7 Sodium (Na) reacts with bromine (Br₂) to form the ionic compound sodium bromide (NaBr). Sodium is in Group 1 of the periodic table. Bromine is in Group 7 of the periodic table.

   a) What names are often given to Groups 1 and 7?
   b) Describe in detail what happens in terms of electrons when sodium reacts with bromine.
   c) Potassium reacts with bromine more vigorously than sodium. Explain why potassium reacts more vigorously than sodium.

8 A yellow solution of bromine water was added dropwise to a colourless solution of sodium iodide. The solution darkened to pale brown.

   a) Explain why the solution darkened.
   b) Write an ionic equation for the reaction that took place.
   c) Explain, in terms of electrons, why this reaction took place.

9 The following four substances are mixed together: salt; water; cyclohexane; diethyl ether. Cyclohexane and diethyl ether are liquids made of organic molecules. Salt is soluble in water but not in cyclohexane or diethyl ether. Cyclohexane and diethyl ether are miscible, but water is not miscible with cyclohexane or diethyl ether. Describe how the four substances could be separated.
### Practice questions

1. How many electrons are there in a potassium ion (K⁺)?
   - A 18
   - B 19
   - C 20
   - D 39 [1 mark]

2. In which of the following atoms is the number of protons greater than the number of neutrons?
   - A ³¹H
   - B ³²He
   - C ³⁹B
   - D ¹⁶O [1 mark]

3. An aluminium atom contains three types of particle.
   a) Copy and complete the table below to show the name, relative mass and relative charge of each particle in an aluminium atom. [4 marks]

<table>
<thead>
<tr>
<th>Particle</th>
<th>Relative charge</th>
<th>Relative mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>1</td>
<td>Very small</td>
</tr>
<tr>
<td>Neutron</td>
<td>0</td>
<td></td>
</tr>
</tbody>
</table>

   b) Complete the sentences below about an aluminium ion by choosing one of the words or numbers in bold. [4 marks]
   i) In an aluminium atom, the protons and neutrons are in the nucleus/shells.
   ii) The number of protons in an aluminium atom is the atomic number/group number/mass number.
   iii) The sum of the number of protons and neutrons in an aluminium atom is the atomic number/group number/mass number.
   iv) The number of electrons in an aluminium atom is 13/14/27.

4. The structure of the atom has caused debate for thousands of years. In the late 19th century the ‘plum pudding model’ of the atom was proposed. This was replaced at the beginning of the 20th century with the nuclear model of the atom which is the basis of the model we use today.
   a) Describe the differences between the ‘plum pudding’ model of the atom and the model of the atom we use today. [5 marks]
   b) The diagram represents an atom of an element. The electrons are missing from the diagram.

   ![Figure 9.33](image)

   11 protons + 12 neutrons.

   ▲ Figure 9.33

5. Mixtures may be separated in the laboratory in many different ways. Three different methods of separating mixtures are shown below.

   ![Figure 9.34](image)

   a) Name each method of separation. [3 marks]
   b) Which method (1, 2 or 3) would be most suitable for obtaining water from potassium chloride solution? [1 mark]
   c) Which method would be most suitable for removing sand from a mixture of sand and water? [1 mark]
   d) What general term is used for liquid A and solid B in method 2? [2 marks]
   e) State why method 2 would not be suitable to separate copper(II) chloride from copper(II) chloride solution. [1 mark]

6. To determine if two different orange drinks X and Y contained the food colourings E102, E101 or E160 a student put a drop of each orange drink and a drop of each food colouring along a pencil line on filter paper.
The filter paper was placed in a tank containing 1 cm depth of solvent. The solvent soaked up the paper and carried different components with it. After 5 minutes, the filter paper was removed and allowed to dry. The results are shown.

a) What is the name of the process used by the student to analyse the two orange drinks? [1 mark]

b) i) Orange drink X contains the food colouring E102. How do the results show this? [1 mark]
   ii) What other food colouring does orange drink X contain? [1 mark]
   iii) Re-draw the diagram and add a spot to show that orange drink Y also contained food colouring E160. [1 mark]
   iv) The line across the bottom of the filter paper was drawn with a pencil not with ink. Why should the line not be drawn with ink? [1 mark]

7 When Group 1 elements react, the atom forms an ion. For example when potassium reacts with water, potassium ions are formed from potassium atoms.

a) Why is potassium stored under oil in the laboratory? [1 mark]

b) Before reacting Group 1 elements with water a risk assessment is carried out. Give two safety precautions, apart from wearing safety glasses, which must be included in the risk assessment for reacting potassium with water. [2 marks]

c) Equal sized pieces of three Group 1 metals are added to separate troughs of water which contain universal indicator. The observations made are recorded in the table. Use the information in the table to answer the questions that follow.

<table>
<thead>
<tr>
<th>Group 1 metal</th>
<th>Observation on reacting with water</th>
<th>Colour of universal indicator solution</th>
</tr>
</thead>
<tbody>
<tr>
<td>Potassium</td>
<td>Melts, Burns with a lilac flame, Moves on the surface of the water, Disappears quickly</td>
<td>Changes colour from green to purple</td>
</tr>
<tr>
<td>Lithium</td>
<td>Floats, Moves on the surface of the water, Eventually disappears</td>
<td>Changes colour from green to purple</td>
</tr>
<tr>
<td>Sodium</td>
<td>Melts, Moves on the surface of the water, Disappears</td>
<td>Changes colour from green to purple</td>
</tr>
</tbody>
</table>

Table 9.26

8 The modern periodic table has been in use for over 100 years. Its development included the work of several chemists including that of Dmitri Mendeleev.

a) Fill in the blanks in the following passage

The modern periodic table arranges the elements in order of increasing atomic ________, whereas early versions of the periodic table arranged them in order of increasing atomic ________. [2 marks]

b) State one other difference between the modern periodic table and Mendeleev’s table. [1 mark]

c) Elements in the periodic table are arranged in groups. The table gives details of some of the groups of the periodic table. Copy and complete the table. [6 marks]

Table 9.27

<table>
<thead>
<tr>
<th>Group number</th>
<th>Name of group</th>
<th>Number of electrons in the outer shell of an atom of this group</th>
<th>Reactive or non-reactive group?</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Table 9.26

7 Many trends in reactivity and physical properties are apparent as a group is descended.

i) State the trend in reactivity as Group 7 is descended. [1 mark]

ii) Name the least reactive element in Group 1. [1 mark]

iii) Astatine is found at the bottom of Group 7. Predict its state at room temperature and pressure. [1 mark]

9 Dmitri Mendeleev produced a table that is the basis of the modern periodic table. Describe the key features of Mendeleev’s table and explain why his table came to be accepted over time by scientists. [6 marks]
Working scientifically: How theories change over time

A version of the periodic table hangs on the wall of almost every chemistry laboratory across the world – it is a powerful icon and a single document that summarises much of our knowledge of chemistry. The history of the periodic table can be traced back over centuries and illustrates how scientific theories change over time.

After the discovery of the new element phosphorus in 1649, scientists began to think about the definition of an element. In 1789 Antoine-Laurent de Lavoisier produced a table similar to that below of simple substances, or elements, which could not be broken down further by chemical reactions.

<table>
<thead>
<tr>
<th>Acid-making elements</th>
<th>Gas making elements</th>
<th>Metallic elements</th>
<th>Earth elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sulfur</td>
<td>Light</td>
<td>Cobalt, mercury, tin</td>
<td>Lime (calcium oxide)</td>
</tr>
<tr>
<td>Phosphorus</td>
<td>Caloric (heat)</td>
<td>Copper, nickel, iron</td>
<td>Magnesia (magnesium oxide)</td>
</tr>
<tr>
<td>Charcoal (carbon)</td>
<td>Oxygen</td>
<td>Gold, lead, silver, zinc</td>
<td>Barytes (barium sulfate)</td>
</tr>
<tr>
<td></td>
<td>Azote (nitrogen)</td>
<td>Manganese, tungsten</td>
<td>Argila (aluminium oxide)</td>
</tr>
<tr>
<td></td>
<td>Hydrogen</td>
<td>Platina (platinum)</td>
<td>Silex (silicon dioxide)</td>
</tr>
</tbody>
</table>

In addition to many elements which form the basis of our modern periodic table, Lavoisier’s list also included ‘light’ and ‘caloric’ (heat) which at the time were believed to be material substances. Lavoisier incorrectly classified some compounds as elements because high temperature smelting equipment or electricity was not available to break down these compounds. The incorrect classification of these compounds as elements was due to a lack of technology as much as a lack of knowledge.
1. What is an element?
2. Which elements in Lavoisier’s table also appear in today’s periodic table?
3. Which group of elements did Lavoisier classify correctly?
4. Why do you think sulfur, phosphorus and charcoal are described as ‘acid-making’ elements?
5. Which substances in Lavoisier’s list, from your own modern knowledge, are compounds? Why do you think Lavoisier thought these were elements?

Following on from the work of Lavoisier, in the early 19th century Johann Döbereiner noted that certain elements could be arranged in groups of three because they have similar properties. For example:

- lithium, sodium and potassium – very reactive metals which produce alkalis with water
- calcium, strontium and barium – reactive metals but with higher melting points and different formulae of their oxides
- chlorine, bromine and iodine – low melting point, coloured, reactive non-metals.

He also noted that the ‘atomic weight’ of the middle element was close to the average of the other two:

\[
\text{Li} = 7; \quad \text{Na} = 23; \quad \text{K} = 39; \quad \frac{7 + 39}{2} = 23
\]

\[
\text{Ca} = 40; \quad \text{Sr} = 88; \quad \text{Ba} = 137; \quad \frac{40 + 137}{2} = 88.5
\]

These groups were called triads and were the first partial representation of a group of elements with similar properties.

6. State the group represented by each of the three Döbereiner triads.
7. Does the final triad listed above follow Döbereiner’s atomic weight rule? Show your working.

You have already learned of the work of Newlands and Mendeleev in the development of the periodic table. It is important to realise that their work was built on the theory and tables suggested by Lavoisier, Döbereiner and others. Use the following questions to think about the ways in which scientific theories and methods develop over time.

8. In what ways is Newland’s periodic table superior to Lavoisier’s classification of the elements?
9. Why is Newland’s classification superior, yet building on Johann Döbereiner’s work?
10. State as many features as you can think of in which the modern periodic table is superior to Mendeleev’s periodic table.

In the 1940s Glenn Seaborg was part of a research team working on ‘nuclear synthesised’ elements with atomic masses beyond the naturally occurring limit of uranium. When isolating the elements americium and curium, he wondered if these elements belonged to a different series which would explain why their chemical properties were different from what was expected. In 1945, against the advice of colleagues, he proposed a significant change to Mendeleev’s table – the actinide series. Today this series is well accepted and included in the periodic table.

Through the history of the periodic table we can easily see:

- the ways in which scientific methods and theories develop over time
- how a variety of concepts and models are used to develop scientific explanations and understanding.
Atoms are so small that we cannot see them. We cannot see them even using the most powerful light microscope because atoms are much smaller than the wavelength of light. However, being able to picture what the particles are like in a substance and how they are bonded to each other is vital to understand chemistry. In this chapter we will examine what the particles are and how they bond together in different substances to help us understand the properties of these different substances.

This chapter covers specification points 5.2.1.1 to 5.2.3.3 and is called Bonding, structure and the properties of matter. It covers ionic, molecular, giant covalent and metallic substances, as well as an overview of types of bonding and structures, nanoscience and the different forms of carbon. Related work on writing formulae and equations can be found in Chapter 14.
Ionic substances

What are ionic substances?
Many substances are made of ions. Ions are electrically charged particles which have a different number of protons (which are positively charged) and electrons (which are negatively charged).

Most compounds made from a combination of metals and non-metals have an ionic structure. For example, sodium chloride is made from sodium (metal) and chlorine (non-metal) and is ionic. Copper sulfate is made from copper (metal), sulfur (non-metal) and oxygen (non-metal) and is also ionic.

The structure of ionic substances
In substances made of ions, there are lots of positive and negative ions in a giant lattice. A giant lattice contains a massive number of particles in a regular structure that continues in all directions throughout the substance (Figure 10.1).
The positive and negative ions are attracted to each other by electrostatic attraction because opposite charges attract. Ionic bonding is the attraction between positive and negative ions. Each ion is attracted to all the ions of opposite charge around it. This attraction is strong and so all ionic substances are solids at room temperature.

Four alternative ways of representing the ions in a lattice are shown in Table 10.1 (using sodium chloride as an example).

### Table 10.1 Methods of representing ions in a lattice.

<table>
<thead>
<tr>
<th>Diagram</th>
<th>Dot and cross diagram</th>
<th>2D space-filling structure</th>
<th>3D space-filling structure</th>
<th>Ball and stick structure</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>[Na$^+$] (2,8)</td>
<td><img src="image" alt="2D structure" /></td>
<td><img src="image" alt="3D structure" /></td>
<td><img src="image" alt="Ball and stick structure" /></td>
</tr>
<tr>
<td></td>
<td>[Cl$^-$] (2,8,8)</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

- **Advantages of this representation**
  - Shows the electronic structure of the ions
  - Very easy to draw
  - Gives very good representation of how the ions are packed together
  - Helps to show how the ions are arranged relative to each other

- **Disadvantages of this representation**
  - Can give the impression that the structure is made of pairs of ions rather than being a continuous structure containing a massive number of ions
  - Can give the impression that the structure is limited to a few ions rather than being a continuous structure with a massive number of ions
  - Only shows the structure in 2D
  - May make you think there are covalent bonds between the ions (there are NO covalent bonds in an ionic lattice)
  - May make you think the ions are a long way apart (but they are packed close together)

**KEY TERMS**

- **Giant lattice** Ionic substances are made up of a giant lattice of positive and negative ions in a regular structure.
- **Ionic bonding** The electrostatic attraction between positive and negative ions.

---

**The properties of ionic substances**

**Melting and boiling points**

In order to melt and boil ionic substances, the strong attraction between the positive and negative ions has to be overcome (Figure 10.2). This is difficult and requires a lot of energy and so ionic substances have high melting and boiling points. For example, sodium chloride melts at 801°C and aluminium oxide melts at 2072°C.

**Electrical conductivity**

An electric current is the flow of electrically charged particles such as ions or electrons. Ionic substances are made of ions, but as a solid the ions cannot move so they cannot conduct electricity. However, when melted, the ions can move and carry charge, so ionic substances will conduct electricity when molten. Many ionic substances dissolve in water and...
will conduct electricity if they dissolve because the ions can move (Figure 10.3).

Figure 10.3 When an ionic substance dissolves in water the ions separate, mix in with the water molecules and move around.

Test yourself

1. What are ions?
2. Explain why ionic substances have high melting points.
3. Explain why ionic substances conduct electricity when molten or dissolved.
4. Explain why ionic substances do not conduct electricity as solids.
5. Ionic substances are made of a giant lattice of ions. What is a giant lattice?
6. Which of the following substances are likely to be ionic: CO₂, PH₃, Fe₂O₃, CH₄O, SiO₂, MgBr₂?

Show you can...

B²⁺ and A⁻ form a compound in which the bonding is ionic.

a) What is the formula of the compound formed between these two ions?
b) Explain fully what is meant by an ionic bond.
c) What type of structure does this compound have? State two of its properties.

The formula of ionic substances

The charge on ions

You can work out the charge on some ions easily. For example, all the elements in

- **Group 1** have one electron in their outer shell and so lose one electron when they react forming 1⁺ ions (e.g. Na⁺, K⁺).
- **Group 2** have two electrons in their outer shell and so lose two electrons when they react forming 2⁺ ions (e.g. Ca²⁺, Mg²⁺).
- **Group 6** have six electrons in their outer shell and so gain two electrons when they react forming 2⁻ ions (e.g. O²⁻, S²⁻).
- **Group 7** have seven electrons in their outer shell and so gain one electron when they react forming 1⁻ ions (e.g. Cl⁻, Br⁻).

These charges and those of other common ions are shown in the Tables 10.2 and 10.3.

<table>
<thead>
<tr>
<th>Group 1 ions (form 1+ ions)</th>
<th>Group 2 ions (form 2+ ions)</th>
<th>Group 3 ions (form 3+ ions)</th>
<th>Others</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li⁺ lithium</td>
<td>Mg²⁺ magnesium</td>
<td>Al³⁺ aluminium</td>
<td>NH₄⁺ ammonium</td>
</tr>
<tr>
<td>Na⁺ sodium</td>
<td>Ca²⁺ calcium</td>
<td>H⁺ hydrogen</td>
<td>Fe²⁺ iron(II)</td>
</tr>
<tr>
<td>K⁺ potassium</td>
<td>Ba²⁺ barium</td>
<td>Cu²⁺ copper(II)</td>
<td>Fe³⁺ iron(III)</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Ag⁺ silver</td>
<td>Pb²⁺ lead</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Zn²⁺ zinc</td>
<td></td>
</tr>
</tbody>
</table>
Table 10.3 Negative ions.

<table>
<thead>
<tr>
<th>Group 6 ions (form 2− ions)</th>
<th>Group 7 ions (form 1− ions)</th>
<th>Others</th>
</tr>
</thead>
<tbody>
<tr>
<td>O2− oxide</td>
<td>F− fluoride</td>
<td>CO3(^{2−}) carbonate</td>
</tr>
<tr>
<td>S2− sulfide</td>
<td>Cl− chloride</td>
<td>OH− hydroxide</td>
</tr>
<tr>
<td></td>
<td>Br− bromide</td>
<td>NO3− nitrate</td>
</tr>
<tr>
<td></td>
<td>I− iodide</td>
<td>SO4(^{2−}) sulfate</td>
</tr>
</tbody>
</table>

What the formula of an ionic substance means

The formula of an ionic substance represents the ratio of the ions in the lattice. For example, in sodium chloride the formula NaCl means that the ratio of sodium (Na\(^{+}\)) ions to (Cl\(^{−}\)) chloride ions in the lattice is 1 : 1. In aluminium oxide, the formula Al₂O₃ means that the ratio of aluminium (Al\(^{3+}\)) ions to oxide (O\(^{2−}\)) ions in the lattice is 2 : 3.

Working out the formula of an ionic substance

In an ionic substance the total number of positive charges must equal the total number of negative charges. This allows us to work out the formula of ionic substances.

Some ions contain atoms of different elements. Examples include sulfate (SO₄\(^{2−}\)), hydroxide (OH\(^{−}\)) and nitrate (NO₃\(^{−}\)). If you need to write more than one of these in a formula, then those ions should be placed in a bracket (Table 10.4).

Table 10.4 Ionic substances and their ions.

<table>
<thead>
<tr>
<th>Name</th>
<th>Positive Ions</th>
<th>Negative Ions</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium chloride</td>
<td>Na(^{+}) (1+ charge)</td>
<td>Cl(^{−}) (1− charge)</td>
<td>NaCl</td>
</tr>
<tr>
<td>Magnesium chloride</td>
<td>Mg(^{2+}) (2+ charges)</td>
<td>Cl(^{−}) (2− charges)</td>
<td>MgCl₂</td>
</tr>
<tr>
<td>Magnesium sulfide</td>
<td>Mg(^{2+}) (2+ charges)</td>
<td>S(^{2−}) (2− charges)</td>
<td>MgS</td>
</tr>
<tr>
<td>Copper(II) sulfate</td>
<td>Cu(^{2+}) (2+ charges)</td>
<td>SO(^{4−}) (2− charges)</td>
<td>CuSO₄</td>
</tr>
<tr>
<td>Sodium carbonate</td>
<td>Na(^{+}) (2+ charges)</td>
<td>CO(^{3−}) (2− charges)</td>
<td>Na₂CO₃</td>
</tr>
<tr>
<td>Ammonium sulfate</td>
<td>NH(^{4+}) (2+ charges)</td>
<td>SO(^{4−}) (2− charges)</td>
<td>(NH₄)₂SO₄</td>
</tr>
<tr>
<td>Calcium nitrate</td>
<td>Ca(^{2+}) (2+ charges)</td>
<td>NO(^{3−}) (2− charges)</td>
<td>Ca(NO₃)₂</td>
</tr>
<tr>
<td>Aluminium oxide</td>
<td>Al(^{3+}) (6+ charges)</td>
<td>O(^{2−}) (6− charges)</td>
<td>Al₂O₃</td>
</tr>
<tr>
<td>Iron(III) hydroxide</td>
<td>Fe(^{3+}) (3+ charges)</td>
<td>OH(^{−}) (3− charges)</td>
<td>Fe(OH)₃</td>
</tr>
</tbody>
</table>

Test yourself

7. What will be the charge on ions of strontium, astatine, selenium and rubidium?
8. What does the formula K₂O mean?
9. Write the formula of the following substances: sodium sulfide, calcium fluoride, magnesium hydroxide, potassium carbonate, barium nitrate, caesium oxide.
10 Bonding, structure and the properties of matter

TIP
When metal atoms lose electrons they form positive ions. When non-metal atoms gain electrons they form negative ions.

TIP
When atoms gain electrons to form ions, the name changes to end in -ide. For example, oxygen atoms gain electrons to form oxide ions, chlorine atoms gain electrons to form chloride ions, sulfur gains electrons to form sulfide ions.

Show you can...
Copy and complete the table to give information about some copper compounds.

Table 10.5

<table>
<thead>
<tr>
<th>Name of compound</th>
<th>Formula of positive ion in compound</th>
<th>Formula of negative ion in compound</th>
<th>Formula of compound</th>
<th>What the numbers in the formula represent</th>
</tr>
</thead>
<tbody>
<tr>
<td>Copper(II) carbonate</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Copper(II) hydroxide</td>
<td></td>
<td>OH⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Copper(II) nitrate</td>
<td>Cu²⁺</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Copper(II) oxide</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Copper(II) sulfate</td>
<td>CuSO₄</td>
<td></td>
<td>1 copper, 1 sulfur, 4 oxygen</td>
<td></td>
</tr>
<tr>
<td>Copper(II) sulfide</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

- The reaction between metals and non-metals

Ionic compounds can be formed when metals react with non-metals. In these reactions, electrons are transferred from the outer shell of the metal atom to the outer shell of the non-metal atom, to produce ions. These ions have the electronic structure of the noble gases (Group 0 elements).

For example, when sodium reacts with chlorine, each sodium atom loses one electron and each chlorine atom gains one electron (Figures 10.4–10.6). This produces sodium ions and chloride ions which have noble gas electronic structures (i.e. sodium chloride).

This can be drawn as a ‘dot and cross’ diagram which only shows the outer shell electrons. Electrons from one atom are shown as dots (●) and electrons from the other atom as crosses (●●) (Figure 10.6).

\[
\text{Na}^{+} + \text{Cl}^{-} \rightarrow [\text{Na}]^{+} + [\text{Cl}]^{-}
\]

\[\begin{align*}
\text{Na} & \quad \text{Cl} \\
(2,8,1) & \quad (2,8,7) \\
\text{Na}^{+} & \quad \text{Cl}^{-} \\
(2,8) & \quad (2,8,8)
\end{align*}\]

Figure 10.5 Sodium reacting with chlorine.

Figure 10.4 Electron transfer during the formation of sodium chloride.

Figure 10.6 Sodium atoms react with chlorine atoms in the ratio 1:1 to form NaCl.
Some other examples of these reactions are shown in Figure 10.7.

(a) Magnesium atoms react with fluorine atoms in the ratio 1:2 to form MgF₂.

(b) Calcium atoms react with oxygen atoms in the ratio 1:1 to form CaO.

(c) Potassium atoms react with sulfur atoms in the ratio 2:1 to form K₂S.

Test yourself

10 Explain why ions are formed when metals react with non-metals.
11 Draw a diagram to show what happens in terms of electrons when lithium reacts with oxygen.

Show you can...

a) Explain fully what happens when magnesium atoms react with oxygen atoms to produce magnesium oxide.
b) Explain fully what happens when magnesium atoms react with fluorine atoms to produce magnesium fluoride.
c) Highlight any similarities and differences between the reaction of magnesium with oxygen and the reaction with fluorine.
What are molecular substances?

Many substances are made of molecules. Molecules are atoms joined together by covalent bonds. A covalent bond is two shared electrons that join atoms together. How covalent bonds form is described later in this chapter.

Many substances are made of molecules. Some non-metal elements are made of molecules. The most common ones are listed in the periodic table as shown in Figure 10.8.

KEY TERMS

Molecule Particle made from atoms joined together by covalent bonds.

Covalent bond Two shared electrons joining atoms together.

Intermolecular forces Weak forces between molecules.

Most compounds are made from a combination of non-metals. For example, water (H₂O) is made from hydrogen (non-metal) and oxygen (non-metal); glucose (C₆H₁₂O₆) is made from carbon (non-metal), hydrogen (non-metal) and oxygen (non-metal).

The structure of molecular substances

A molecular substance is made of many identical molecules that are not joined to each other (Figure 10.9). Within each molecule, the atoms are joined together by the very strong covalent bonds. However, the molecules are not bonded to each other. There are only some weak forces between the molecules – these weak forces are intermolecular forces.

Molecules are often quite small, containing just a few atoms, but some substances are made of big molecules (e.g. wax and many polymers).

The properties of molecular substances

Melting and boiling points

Molecules are not bonded to each other. The intermolecular forces (forces between molecules) are only weak and so are easy to overcome. This means that molecular substances have low melting and boiling points. Many molecular substances with small molecules are gases and liquids at room temperature. For example, methane boils at -162°C and water boils at 100°C.
Generally, the bigger the molecules, the stronger the forces between the molecules and so the higher the melting and boiling points. Molecules of glucose are quite large and it melts at 146°C.

When molecular substances change state, the covalent bonds do not break. For example, water molecules are identical as H₂O whether it is steam, water or ice (Figure 10.10). No covalent bonds are broken when water changes state.

<table>
<thead>
<tr>
<th>Solid (Ice)</th>
<th>Gas (steam)</th>
</tr>
</thead>
<tbody>
<tr>
<td><img src="image1" alt="Solid structure" /></td>
<td><img src="image2" alt="Gas structure" /></td>
</tr>
</tbody>
</table>

**Figure 10.10** The structure of water, a molecular substance as a solid, liquid and gas. No covalent bonds are broken when it changes state.

**Electrical conductivity**

Molecules are electrically neutral which means that molecular substances do not conduct electricity at all. They do not contain delocalised electrons or charged ions.

**Test yourself**

12 What are molecules?
13 What is a covalent bond?
14 Explain why molecular substances have low melting and boiling points.
15 What happens to the covalent bonds in a molecular substance when it melts and boils?
16 Explain why molecular substances do not conduct electricity.
17 Which of the following substances are likely to be molecular: H₂S, Na₂O, KNO₃, ZnBr₂, CO, N₂H₄, C₂H₆O?

**Show you can...**

Metal oxides and non-metal oxides have different properties. Sulfur dioxide, a non-metal oxide has a melting point of −72°C and calcium oxide, a metal oxide, has a melting point of 2613°C. Explain why the melting point of sulfur dioxide is low but that of calcium oxide is high.

○ **The formula of molecular substances**

Molecular substances have two formulae, the empirical formula and the molecular formula. The molecular formula is the one that is normally used.
10 Bonding, structure and the properties of matter

Table 10.6 Molecular and empirical formulae of molecules.

<table>
<thead>
<tr>
<th>Diagram of molecule</th>
<th>Molecular formula (gives the number of atoms of each element in each molecule)</th>
<th>Empirical formula (gives the simplest ratio of the atoms of each element in the substance)</th>
</tr>
</thead>
</table>
| Butane              | C₄H₁₀  
There are 4 C atoms and 10 H atoms in each molecule | C₂H₅  
Simplest ratio of C:H = 2:5 |
| Water               | H₂O  
There are 2 H atoms and 1 O atom in each molecule | H₂O  
Simplest ratio of H:O = 2:1 |
| Glucose             | C₆H₁₂O₆  
There are 6 C atoms, 12 H atoms and 6 O atoms in each molecule | CH₂O  
Simplest ratio of C:H:O = 1:2:1 |

Test yourself

18 Benzene has the molecular formula C₆H₆. What does this tell us about benzene?
19 What is the empirical formula of benzene? What does this tell us about benzene?

Show you can...

a) Copy and complete the table:

<table>
<thead>
<tr>
<th>Molecular formula</th>
<th>Empirical formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>C₂H₆</td>
<td></td>
</tr>
<tr>
<td>C₃H₇N₂O₂</td>
<td></td>
</tr>
<tr>
<td>C₂H₁₀O₂</td>
<td></td>
</tr>
</tbody>
</table>

b) Are the empirical formula and the molecular formula of a substance always different? Using an example, explain your answer.

Drawing molecules

When atoms join together to form molecules they share electrons in order to obtain noble gas electronic structures. A covalent bond is two shared electrons joining atoms together. Table 10.8 shows how many covalent bonds atoms typically form.

Table 10.8 The relationship between number of covalent bonds and electronic structure.

<table>
<thead>
<tr>
<th>Atoms</th>
<th>H</th>
<th>Group 4 atoms (e.g. C, Si)</th>
<th>Group 5 atoms (e.g. N, P)</th>
<th>Group 6 atoms (e.g. O, S)</th>
<th>Group 7 atoms (e.g. F, Cl, Br, I)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Number of electrons in their outer shell</td>
<td>1</td>
<td>4</td>
<td>5</td>
<td>6</td>
<td>7</td>
</tr>
<tr>
<td>Number of electrons needed to obtain a noble gas electronic structure</td>
<td>1</td>
<td>4</td>
<td>3</td>
<td>2</td>
<td>1</td>
</tr>
<tr>
<td>Number of covalent bonds formed</td>
<td>1</td>
<td>4</td>
<td>3</td>
<td>2</td>
<td>1</td>
</tr>
</tbody>
</table>
There are several ways to show how atoms join together by sharing electrons in covalent bonds to form molecules (Table 10.9).

### Table 10.9  Different ways of representing covalent bonding.

<table>
<thead>
<tr>
<th>Dot and cross diagram showing all electrons and shell circles</th>
<th>Dot and cross diagram showing only outer shell electrons and shell circles</th>
<th>Dot and cross diagram showing only outer shell electrons</th>
<th>Stick diagram</th>
<th>Ball and stick diagram</th>
<th>Space-filling diagram</th>
</tr>
</thead>
<tbody>
<tr>
<td><img src="image1.png" alt="Diagram" /></td>
<td><img src="image2.png" alt="Diagram" /></td>
<td><img src="image3.png" alt="Diagram" /></td>
<td><img src="image4.png" alt="Diagram" /></td>
<td><img src="image5.png" alt="Diagram" /></td>
<td><img src="image6.png" alt="Diagram" /></td>
</tr>
</tbody>
</table>

Note that the \(\text{●}\) and \(\text{●}\) represent electrons that came from different atoms.

Each stick (or line) represents one covalent bond (i.e. 2 shared electrons).

A good representation of a molecule showing how atoms merge into each other, but the covalent bonds are not visible.

---

When drawing stick diagrams, each atom makes the number of covalent bonds shown in Table 10.8. When drawing dot-cross diagrams, each single covalent bond is made up of two electrons. Atoms can make **double** and **triple** covalent bonds (Figure 10.11). A double covalent bond contains four electrons (two from each atom), while a triple bond contains six electrons (three from each atom).

Any outer shell electrons that are not used up in making covalent bonds are found in non-bonding electron pairs, often called lone pairs.

---

**Test yourself**

1. What does each stick represent in a stick diagram?
2. What do the dots and crosses represent in dot and cross diagrams?
3. Draw a stick diagram and a dot and cross diagram for \(\text{H}_2\text{S}\).
4. Draw a stick diagram and a dot and cross diagram for \(\text{CS}_2\).

---

**Show you can...**

Phosphorus bonds with hydrogen to form phosphine. \(\text{PH}_3\) is a colourless gas which has an unpleasant, rotting fish odour. Phosphorus also bonds with chlorine to form phosphorus trichloride which is a toxic colourless liquid.

a) Draw a dot and cross diagram to show the bonding in \(\text{PH}_3\).
b) Suggest the formula of phosphorus trichloride.
c) Draw a dot and cross diagram to show the bonding in phosphorus trichloride.
d) Using your diagram from c), explain what is meant by a covalent bond and a non-bonding electron pair.
Polymers

There are many different types of polymer (plastics), including polythene, PVC, Perspex, Teflon and polystyrene. Polymers contain very large molecules, often with hundreds or thousands or atoms. Within each molecule, the atoms are joined to each other by strong covalent bonds and so are solids at room temperature (Figure 10.12). The basic sub-units of polymers are called monomers.

Polymers can be represented in the form:

\[
\begin{align*}
\text{polythene in made from many sub-units (monomers) of ethene} \\
\text{e.g. polythene} \\
\text{the point where other units (monomers) join} \\
\text{the number of times the basic structure in brackets is repeated}
\end{align*}
\]

KEY TERMS

**Polymer** Long chain molecule made from joining lots of small molecules together by covalent bonds.

**Monomer** The building block (molecule) of a polymer.

**TIP**

Polymers are sometimes called macromolecules as they have very long molecules.
Testing the electrical conductivity of ionic and molecular covalent substances

To investigate the conduction of electricity by a number of compounds in aqueous solution.

The apparatus was set up as shown in the diagram.

Questions
1. Describe the experimental method which you would use to test the solutions using the apparatus shown.
2. Copy and complete the results table.
3. Using the results from column three and four of the table write a conclusion for this experiment stating and explaining any trends shown in the results.
4. Would the results be different if solid copper(II) sulfate was used in place of copper(II) sulfate solution? Explain your answer.
5. Predict and explain the results you would obtain for calcium nitrate solution.
6. Predict and explain the results you would obtain for bromine solution.

Table 10.10

<table>
<thead>
<tr>
<th>Test solution</th>
<th>Does the bulb light?</th>
<th>Does the substance conduct electricity?</th>
<th>Does the substance contain ionic or covalent bonding?</th>
</tr>
</thead>
<tbody>
<tr>
<td>Copper(II) sulfate</td>
<td>yes</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ethanol (C₂H₅OH)</td>
<td>no</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Magnesium sulfate</td>
<td>yes</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Potassium iodide</td>
<td>yes</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Glucose (C₆H₁₂O₆)</td>
<td>no</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Sodium chloride</td>
<td>yes</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Giant covalent substances

What are giant covalent substances?

There are a few substances that have atoms joined by covalent bonds in a continuous network. Common examples are

- diamond, C (a form of carbon) – studied in detail later in the chapter ('The different forms of carbon' section)
- graphite, C (a form of carbon) – studied in detail later in the chapter ('The different forms of carbon' section)
- silicon, Si
- silicon dioxide, SiO₂ (also known as silica).

The structure of giant covalent substances

In a giant covalent substance all the atoms are in a giant lattice. They are all joined together by covalent bonds in a continuous network throughout the structure (Figure 10.14).

These substances are not molecules. In molecular substances, there are lots of separate molecules with the atoms in each molecule joined by covalent bonds but the molecules are not joined together.
The properties of giant covalent substances

Melting and boiling points
In order to melt a giant covalent substance, many covalent bonds have to be broken. Covalent bonds are very strong and so it takes a lot of energy to break them. Therefore, giant covalent substances are solids with very high melting and boiling points. For example, diamond melts at over 3500°C.

Electrical conductivity
Most giant covalent substances do not conduct electricity because they do not contain any delocalised electrons. However, graphite does as it does have some delocalised electrons. Delocalised electrons are able to move throughout the substance.

Test yourself
24 Describe the structure of a giant covalent substance.
25 Why do giant covalent substances have very high melting points?
26 Why do giant covalent substances, except graphite, not conduct electricity?

Show you can...
Carbon and silicon both form dioxides; carbon dioxide and silicon dioxide. Carbon dioxide is a gas at room temperature and silicon dioxide is a solid with melting point of 1610°C.
a) Copy and complete the table:

<table>
<thead>
<tr>
<th>Substance</th>
<th>Type of bonding</th>
<th>Structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon dioxide</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Silicon dioxide</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

b) Explain why silicon dioxide is a solid with a high melting point and carbon dioxide is a gas with a low melting point.

Metallic substances

What are metallic substances?
Metals are metallic substances. Over three-quarters of all the elements are metals and have a metallic structure.

The structure of metallic substances
Metals consist of a giant lattice of positively charged atoms arranged in a regular pattern. The outer shell electrons from each atom are delocalised which means they are free to move throughout the whole structure (Figures 10.15 and 10.16).
There is a strong attraction between the positive nucleus of these atoms and the delocalised electrons. This attraction between the nucleus and the delocalised electrons is called **metallic bonding**.

**The properties of metallic substances**

**Melting and boiling points**

In metals, the metallic bonding is strong. This means that most metals are solids and have **high melting and boiling points**. For example, aluminium melts at 660°C and iron melts at 1538°C.

**Electrical conductivity**

Metals are good conductors of electricity because the delocalised electrons are able to move through the structure and carry electrical charge through the metal (Figure 10.17).

**Thermal conductivity**

Metals are also good thermal conductors. The thermal energy is transferred by the delocalised electrons.

**Malleability**

Metals are **malleable**, which means they can be bent and hammered into shape. This is because the layers of atoms can slide over each other while maintaining the metallic bonding (Figure 10.18). This makes metals soft.

**KEY TERMS**

- **Metallic bonding**: The attraction between the nucleus of metal atoms and delocalised electrons.
- **Malleable**: Can be hammered into shape.
Alloys

Pure metals are very malleable. This can make them too soft for most uses as they lose their shape easily. Metals can be made more useful by making them into alloys.

An alloy is a mixture of a metal with small amounts of other elements, usually other metals. Pure metals such as aluminium, iron, copper and gold are rarely used, and alloys of these metals are used instead. For example, steels are alloys made from iron. Alloys of gold are used for making jewellery as pure gold would lose its shape too easily.

Alloys also have metallic structures. However, some of the atoms in the alloy are a different size to those of the metal. This distorts the layers in the structure and makes it much more difficult for the layers of atoms to slide over each other (Figure 10.19).

Test yourself
27 What is metallic bonding?
28 Why do metals have high melting points?
29 Why do metals conduct electricity?
30 Why are metals malleable?
31 What are alloys?
32 Why are alloys harder than pure metals?

Show you can...

Aluminium is used in overhead electricity cables and to make saucepans.

a) Explain why aluminium is a good conductor of electricity.
b) State two other reasons why aluminium is used in overhead electricity cables.
c) Explain in terms of structure why aluminium is malleable.
d) Aluminium oxide is a compound of aluminium. Compare and contrast how aluminium and aluminium oxide conduct electricity.

The physical properties of Group 1 elements

The table shows some physical properties of Group 1 elements.

Table 10.12

<table>
<thead>
<tr>
<th>Element</th>
<th>Melting point in °C</th>
<th>Boiling point in °C</th>
<th>Density in g/cm³</th>
<th>Electrical conductivity</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td>180</td>
<td>1340</td>
<td>0.53</td>
<td>Good</td>
</tr>
<tr>
<td>Na</td>
<td>98</td>
<td>880</td>
<td>0.97</td>
<td>Good</td>
</tr>
<tr>
<td>K</td>
<td>63</td>
<td>766</td>
<td>0.89</td>
<td>Good</td>
</tr>
<tr>
<td>Rb</td>
<td>39</td>
<td>686</td>
<td>1.53</td>
<td>Good</td>
</tr>
<tr>
<td>Cs</td>
<td>28</td>
<td>669</td>
<td>1.93</td>
<td>Good</td>
</tr>
</tbody>
</table>

Questions
1 Use the table to state a property of Group 1 metals which is common to all metals.
2 Describe the type of bonding found in Group 1 metals.
3 What is meant by the term melting?
Overview of types of bonding and structures

○ Types of bonding
There are three types of bonding that are summarised in the Table 10.13.

Table 10.13 The three types of bonding.

<table>
<thead>
<tr>
<th>Description</th>
<th>Ionic bonding</th>
<th>Covalent bonding</th>
<th>Metallic bonding</th>
</tr>
</thead>
<tbody>
<tr>
<td>The electrostatic attraction between positive and negative ions</td>
<td>Atoms that are joined together by sharing pairs of electrons</td>
<td>The attraction between positive nucleus of metal atoms and delocalised outer shell electrons</td>
<td></td>
</tr>
<tr>
<td>Which substances have this bonding</td>
<td>Ionic compounds</td>
<td>Molecular substances</td>
<td>Metallic substances</td>
</tr>
<tr>
<td></td>
<td>Giant covalent substances</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

○ Types of structure
There are five types of structure that are summarised in the Table 10.14. The elements in Group 0 of the periodic table (the noble gases) have a unique structure called monatomic which is very similar to molecules except that the particles are individual atoms and not molecules. The elements in Group 0 are stable and have no need to lose, gain or share electrons.

Table 10.14 Different types of structure.

<table>
<thead>
<tr>
<th>Which substances have this structure</th>
<th>Monatomic</th>
<th>Molecular</th>
<th>Giant covalent</th>
<th>Ionic</th>
<th>Metallic</th>
</tr>
</thead>
<tbody>
<tr>
<td>Made up of many separate atoms. The atoms are not bonded to each other. There are very weak forces of attraction between the atoms</td>
<td>Group 0 elements</td>
<td>Many non-metal elements</td>
<td>Diamond</td>
<td>Most compounds made from a combination of metals with non-metals</td>
<td>Metals and alloys</td>
</tr>
<tr>
<td>Made up of many separate molecules. The atoms within each molecule are joined by covalent bonds. There are no bonds between molecules. There are only weak forces of attraction between the molecules</td>
<td>Diamond</td>
<td>Graphite</td>
<td>Silicon dioxide</td>
<td>Made of a giant lattice of atoms joined to each other by covalent bonds</td>
<td>Made of a giant lattice of positive and negative ions. There are strong electrostatic forces of attraction between the positive and negative ions</td>
</tr>
<tr>
<td>Made of a giant lattice of atoms joined to each other by covalent bonds</td>
<td>Made of a giant lattice of positive and negative ions. There are strong electrostatic forces of attraction between the positive and negative ions</td>
<td>Made of a giant lattice of metal atoms with a cloud of delocalised outer shell electrons. There are strong forces of attraction between the positive nucleus of the metal atoms and the delocalised electrons</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
### Melting and boiling points

<table>
<thead>
<tr>
<th>Bonding Type</th>
<th>Monatomic</th>
<th>Molecular</th>
<th>Giant covalent</th>
<th>Ionic</th>
<th>Metallic</th>
</tr>
</thead>
<tbody>
<tr>
<td>Melting point</td>
<td>VERY LOW as it requires little energy to overcome the very weak forces between atoms</td>
<td>LOW as it requires little energy to overcome the weak forces between molecules</td>
<td>VERY HIGH as it requires a lot of energy to break lots of strong covalent bonds</td>
<td>HIGH as it requires a lot of energy to overcome the strong attraction between the positive and negative ions</td>
<td>HIGH as it requires a lot of energy to overcome the strong attraction between the positive nucleus of the metal atoms and delocalised electrons</td>
</tr>
<tr>
<td>Boiling point</td>
<td>VERY LOW as it requires little energy to overcome the very weak forces between atoms</td>
<td>LOW as it requires little energy to overcome the weak forces between molecules</td>
<td>VERY HIGH as it requires a lot of energy to break lots of strong covalent bonds</td>
<td>HIGH as it requires a lot of energy to overcome the strong attraction between the positive and negative ions</td>
<td>HIGH as it requires a lot of energy to overcome the strong attraction between the positive nucleus of the metal atoms and delocalised electrons</td>
</tr>
</tbody>
</table>

### Electrical conductivity

<table>
<thead>
<tr>
<th>Bonding Type</th>
<th>Monatomic</th>
<th>Molecular</th>
<th>Giant covalent</th>
<th>Ionic</th>
<th>Metallic</th>
</tr>
</thead>
<tbody>
<tr>
<td>Conduction</td>
<td>NON-CONDUCTOR as atoms are neutral and there are no delocalised electrons or ions</td>
<td>NON-CONDUCTOR as molecules are neutral and there are no delocalised electrons or ions</td>
<td>NON-CONDUCTOR (except graphite) as there are no delocalised electrons (graphite does have delocalised electrons)</td>
<td>Solid = NON-CONDUCTOR as ions cannot move</td>
<td>CONDUCTOR as outer shell electrons are delocalised and can carry the charge</td>
</tr>
</tbody>
</table>

### Test yourself

33 For each of the following substances, state the type of bonding and the type of structure it is likely to have:
- a) copper(II) oxide (CuO)
- b) diamond (C)
- c) lead carbonate (PbCO₃)
- d) phosphorus oxide (P₄O₁₀)
- e) argon (Ar)
- f) copper (Cu)

### Show you can...

Substances may be classified in terms of their physical properties. Use the table to answer the following questions:

**Table 10.15**

<table>
<thead>
<tr>
<th>Substance</th>
<th>Melting point in °C</th>
<th>Boiling point in °C</th>
<th>Electrical conductivity as solid</th>
<th>Electrical conductivity as liquid</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>3720</td>
<td>4827</td>
<td>Good</td>
<td>Poor</td>
</tr>
<tr>
<td>B</td>
<td>−95</td>
<td>69</td>
<td>Poor</td>
<td>Poor</td>
</tr>
<tr>
<td>C</td>
<td>327</td>
<td>1760</td>
<td>Good</td>
<td>Good</td>
</tr>
<tr>
<td>D</td>
<td>3550</td>
<td>4827</td>
<td>Poor</td>
<td>Good</td>
</tr>
<tr>
<td>E</td>
<td>801</td>
<td>1413</td>
<td>Poor</td>
<td>Good</td>
</tr>
</tbody>
</table>

a) Which substance could be sodium chloride? Explain your answer.
b) Which substance consists of small covalent molecules? Explain your answer.
c) Explain why substance A could not be diamond.
d) Which substance is a metal?

### States of matter

The three **states of matter** are solid, liquid and gas (Figure 10.20). Substances change state at their melting and boiling points. A substance is a:
- solid at temperatures below its melting point
- liquid at temperatures between its melting and boiling point
- gas at temperatures above its boiling point.
The amount of energy needed for substances to melt and boil depends on the strength of the forces or bonds between their particles. The stronger the forces or bonds between the particles, the higher their melting and boiling points. For example, giant covalent substances have very high melting and boiling points as strong covalent bonds have to be broken. Molecular substances have low melting and boiling points as there are only weak forces between the molecules that are easy to overcome.

You may have used a very simple model to represent the particles in solids, liquids and gases like the top row in Figure 10.21. It can help to understand how particles are arranged when substances are in each state.

The top row in Figure 10.21 is particularly over-simplistic. The particles in most substances are actually molecules or ions and they are not solid spheres. These diagrams also do not show that there are forces or bonds between the particles. They also cannot show whether the particles are moving around or not.

**Test yourself**

34 Use the data in the table to answer the questions that follow.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Melting point in °C</th>
<th>Boiling point in °C</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>45</td>
<td>137</td>
</tr>
<tr>
<td>B</td>
<td>595</td>
<td>984</td>
</tr>
<tr>
<td>C</td>
<td>−30</td>
<td>56</td>
</tr>
<tr>
<td>D</td>
<td>−189</td>
<td>−186</td>
</tr>
<tr>
<td>E</td>
<td>186</td>
<td>302</td>
</tr>
</tbody>
</table>

a) Which substance(s) is/are gases at room temperature?
b) Which substance(s) is/are liquids at 100°C?
c) Which substance(s) is/are solids at 100°C?
d) Which substance is a liquid over the widest temperature range?
10 Bonding, structure and the properties of matter

The melting points and boiling points of six substances are shown in the table.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Melting point in °C</th>
<th>Boiling point in °C</th>
<th>Type of bonding present</th>
</tr>
</thead>
<tbody>
<tr>
<td>N₂</td>
<td>−210</td>
<td>−196</td>
<td></td>
</tr>
<tr>
<td>CS₂</td>
<td>−112</td>
<td>46</td>
<td></td>
</tr>
<tr>
<td>NH₃</td>
<td>−78</td>
<td>−34</td>
<td></td>
</tr>
<tr>
<td>Br₂</td>
<td>−7</td>
<td>59</td>
<td>covalent</td>
</tr>
<tr>
<td>LiCl</td>
<td>605</td>
<td>1137</td>
<td></td>
</tr>
<tr>
<td>Cu</td>
<td>1084</td>
<td>2562</td>
<td></td>
</tr>
</tbody>
</table>

a) Copy and complete the table by deciding if the bonding in each substance is ionic, covalent or metallic.

b) i) Which element is a solid at room temperature?
   ii) Which compound is a liquid at room temperature?
   iii) Which compound is a gas at room temperature?
   iv) Which element will condense when cooled to room temperature from 100°C?
   v) Which compound will freeze first on cooling from room temperature to a very low temperature?

The different forms of carbon

There are several different forms of the element carbon which usefully illustrate several parts of this chapter.

- **Diamond**
  
  Diamond is probably the best known form of carbon. It has a giant covalent structure with all the carbons joined by covalent bonds in a giant lattice. This can be thought of as a continuous network of atoms linked by covalent bonds (Figure 10.22).

   ![Figure 10.22 Diamond has a giant covalent structure with carbon atoms joined by covalent bonds in a giant lattice.](image)

   Close examination of the structure of diamond shows that each carbon atom is covalently bonded to four other carbon atoms. Figure 10.23 shows a very small part of the diamond lattice to help see how the carbon atoms are bonded together.

   ![Figure 10.23 A tiny part of the diamond lattice. Each C atom is bonded to four others.](image)
Graphite

Graphite is another form of carbon. It is the grey substance that runs through the inside of a pencil and rubs off onto the paper. Like diamond, graphite has a giant covalent structure with all the carbons joined by covalent bonds in a giant lattice (Figure 10.24). However, the carbon atoms are bonded together in flat layers. The layers of atoms are not bonded to each other and there are only very weak attractive forces between these layers of atoms.

Close examination of the structure of graphite shows that each carbon atom is covalently bonded to three other carbon atoms (Figure 10.25).

The diagram shows a very small part of the graphite lattice to help see how the carbon atoms are bonded together.

This bonding leaves one outer shell electron on each carbon atom that is not used in bonding. These electrons become delocalised and are free to move along the layers.

Table 10.18 Comparison of the physical properties of diamond and graphite.

<table>
<thead>
<tr>
<th></th>
<th>Diamond</th>
<th>Graphite</th>
</tr>
</thead>
<tbody>
<tr>
<td>Melting point</td>
<td>Very high melting point (over 3500°C) because lots of strong covalent bonds need to be broken</td>
<td>Very high melting point (over 3500°C) because lots of strong covalent bonds need to be broken</td>
</tr>
<tr>
<td>Hardness</td>
<td>Very hard (the hardest natural substance) because the atoms are arranged in a very rigid continuous network held together by strong covalent bonds</td>
<td>Soft because the layers of atoms are not bonded together and so can easily slide over each other</td>
</tr>
<tr>
<td>Conductivity</td>
<td>Does not conduct because it contains no delocalised electrons</td>
<td>Conducts because it contains delocalised electrons (one from each atom) that move along the layers and carry charge through the graphite</td>
</tr>
<tr>
<td>Uses</td>
<td>Drill and saw tips – due to its hardness</td>
<td>Electrodes – as it conducts electricity In pencils – it rubs off easily onto the paper because layers can easily move over each other</td>
</tr>
</tbody>
</table>

Graphene

Graphene is a new substance. It is a single layer of graphite (Figure 10.26). Scientists at the University of Manchester won a Nobel Prize in 2010 for their work on graphene.

Graphene has some remarkable properties. It is extremely thin being just one atom thick, but is extremely strong due to its giant covalent structure. It is also semi see-through as it is so thin. In a similar way to graphite, it is a thermal and electrical conductor due to having some delocalised electrons.
Graphene is a very exciting new material and a lot of research is being done to make use of it. Its properties make it very useful in electronics (e.g. touchscreens) and composite materials (e.g. carbon fibres).

**Fullerenes**

The molecule C\textsubscript{60} was identified in the 1980s as another form of carbon. The molecule has a shape that resembles a football (Figure 10.27). It was named buckminsterfullerene after the American architect Richard Buckminster Fuller who built domes that had similar structures. C\textsubscript{60} is often referred to as a buckyball and was the first fullerene produced. Scientists at the University of Sussex won a Nobel Prize in 1996 for their work on fullerenes.

![Figure 10.27 C\textsubscript{60} has the shape that resembles a football.](image)

A whole family of similar molecules, such as C\textsubscript{70} and C\textsubscript{84}, have been produced. These molecules are all called **fullerenes**. The structure of fullerenes is based on carbon atoms in **hexagonal rings** (rings of six atoms), but some rings have five or seven atoms. They all have a hollow part in the centre of the molecule.

**Fullerenes are being used:**
- For **delivery of drugs** into specific parts of the body and/or cells – the drugs are often carried inside the hollow centre of the fullerene molecule.
- In **lubricants** to reduce friction when metal parts of machines move past each other – the spherical shape of the molecule allows molecules to roll past each other.
- As **catalysts** – a lot of research is taking place into the use of fullerenes as catalysts and a wide range of potential applications are being found.

**Carbon nanotubes**

Carbon nanotubes are **cylindrical fullerenes**, sometimes called buckytubes (Figure 10.28). They have very high length to diameter ratios, significantly higher than for any other material. They can also be thought of as being tubes of graphene sheets.

![Figure 10.28 A carbon nanotube.](image)
These carbon nanotubes have some excellent properties making them very useful. They have:

- **High tensile strength** – in other words it is very strong when it is pulled – this is due to the many strong covalent bonds throughout its structure.
- **High thermal and electrical conductance** – this is due to some of the electrons being delocalised.

Carbon nanotubes have many uses, for example to reinforce the materials used to make sports equipment like tennis racquets (Figure 10.29) and golf clubs.

![Figure 10.29 Carbon nanotubes are very strong and are used to reinforce tennis racquets.](image.png)

### Test yourself

35 Explain why diamond and graphite both have very high melting points.

36 Explain why diamond is hard but graphite is soft.

37 Explain why graphite conducts electricity but diamond does not.

38 What are fullerenes?

39 Carbon nanotubes are used to reinforce and strengthen tennis racquets. Explain why carbon nanotubes strengthen materials.

40 A typical carbon nanotube is 12 cm long and has a diameter of 1 nm. Calculate the length to diameter ratio of this carbon nanotube.

### Show you can...

Copy and complete the table to give information about the different structures and uses of carbon.

<table>
<thead>
<tr>
<th>Table 10.19</th>
<th>Graphite</th>
<th>Diamond</th>
<th>Graphene</th>
<th>fullereneS</th>
<th>Carbon nanotubes</th>
</tr>
</thead>
<tbody>
<tr>
<td>Description of structure</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Example of use</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Chapter review questions

1. Four structure types are: ionic, metallic, molecular and giant covalent.
   a) Which of these structure types usually have low melting and boiling points?
   b) Which of these structure types usually conduct electricity as solids?
   c) Which of these structure types usually conduct electricity when melted?

2. Carbon dioxide is a molecular compound with the formula CO₂.
   a) Explain why carbon dioxide has a low boiling point.
   b) Explain why carbon dioxide does not conduct electricity.

3. Iron is a metal with the formula Fe.
   a) Explain why iron has a high melting point.
   b) Explain why iron conducts electricity.
   c) Explain why pure iron is soft.
   d) Steels are alloys of iron. Explain why steels are harder than pure iron.

4. Potassium fluoride is an ionic compound containing potassium (K⁺) ions and fluoride (F⁻) ions.
   a) Give the electronic structure of potassium (K⁺) ions.
   b) Give the electronic structure of fluoride (F⁻) ions.
   c) Give the formula of potassium fluoride.
   d) Explain why potassium fluoride has a high melting point.
   e) Explain why potassium fluoride does not conduct electricity as a solid.
   f) Explain why potassium fluoride conducts electricity when molten or dissolved.

5. Decide whether each of the following substances has an ionic, molecular, giant covalent or metallic structure.
   a) zinc (Zn)
   b) ethane (C₂H₆)
   c) diamond (C)
   d) magnesium oxide (MgO)
   e) iodine trifluoride (IF₃)
   f) potassium carbonate (K₂CO₃)

6. Decide whether each of the following substances has an ionic, molecular, giant covalent or metallic structure.

Table 10.20

<table>
<thead>
<tr>
<th>Substance</th>
<th>Melting point in °C</th>
<th>Boiling point in °C</th>
<th>Electrical conductivity as solid</th>
<th>Electrical conductivity as liquid</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>838</td>
<td>1239</td>
<td>Does not conduct</td>
<td>Conducts</td>
</tr>
<tr>
<td>B</td>
<td>89</td>
<td>236</td>
<td>Does not conduct</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>C</td>
<td>678</td>
<td>935</td>
<td>Conducts</td>
<td>Conducts</td>
</tr>
<tr>
<td>D</td>
<td>1056</td>
<td>1438</td>
<td>Does not conduct</td>
<td>Conducts</td>
</tr>
<tr>
<td>E</td>
<td>2850</td>
<td>3850</td>
<td>Does not conduct</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>F</td>
<td>−39</td>
<td>357</td>
<td>Conducts</td>
<td>Conducts</td>
</tr>
</tbody>
</table>
7 Draw a stick and a dot-cross diagram for each of the following molecules.
   a) phosphine (PH₃)
   b) bromine (Br₂)
   c) carbon dioxide (CO₂)

8 Calcium reacts with chlorine to form the ionic compound calcium chloride. Draw diagrams to show the electronic structure in calcium atoms, chlorine atoms, calcium ions and chloride ions.

9 Work out the formula of the following ionic compounds. The charge of some ions is given (sulfate = SO₄²⁻, hydroxide = OH⁻)
   a) potassium oxide
   b) magnesium fluoride
   c) lithium sulfide
   d) iron(III) sulfate
   e) copper(II) hydroxide

10 Ethyne is a molecule with the molecular formula C₂H₂.
   a) Explain what this formula tells us about ethyne.
   b) Draw a stick diagram to show the covalent bonds in a molecule of ethyne.
   c) Draw a dot-cross diagram to show the outer shell electrons in a molecule of ethyne.

11 The diagram shows part of a carbon nanotube. They are used to reinforce the materials used to make tennis racquets as they have high tensile strength.

   ▲ Figure 10.30
   a) How many covalent bonds does each carbon atom make in a carbon nanotube?
   b) Explain clearly by considering your answer to (a) why nanotubes can conduct electricity.
   c) Explain clearly why carbon nanotubes have high tensile strength.
Practice questions

1 The elements P and Q are in found in Groups 2 and 7 respectively, of the periodic table. Which one of the following shows the formula and the bonding type of the compound that they form? [1 mark]
   A \( \text{P}_2\text{Q}_6 \) covalent
   B \( \text{P}_2\text{Q}_6 \) ionic
   C \( \text{P}_2\text{Q}_4 \) covalent
   D \( \text{P}_2\text{Q}_4 \) ionic

2 Which one of the following does not have a giant covalent structure? [1 mark]
   A diamond
   B graphene
   C graphite
   D sulfur dioxide

3 A dot and cross diagram is given here:

![Figure 10.31]

a) Name and write the formula for this compound. [2 marks]

b) On a copy of the diagram above, use an arrow to label:
   i) A covalent bond. [2 marks]
   ii) A non-bonded pair of electrons. [2 marks]

c) Using a line to represent a single covalent bond, redraw the diagram shown above. [1 mark]

d) What is meant by the term single covalent bond? [2 marks]

4 The following diagram shows some changes between the states of matter.

![Figure 10.32]

a) What is the name for each of the changes labelled A, B and C? [3 marks]

b) Which of the changes A, B, or C is achieved by a decrease in temperature? [1 mark]

c) Draw a diagram to represent the arrangement of atoms in a solid using small solid spheres to represent atoms. [1 mark]

d) Explain the limitations of this simple particle theory in relation to change in state. [2 marks]

e) Explain why the melting points of some solids are greater than others. [2 marks]

5 The diagrams show two different structures of the Group 4 element carbon and two compounds of Group 4 elements.

![Figure 10.33]

a) Which two of the diagrams represent different structures of the element carbon? [1 mark]

b) The substances are carbon dioxide, diamond, graphite and silicon dioxide. Name each substance, A to D. [4 marks]

c) Name the type of bonding which occurs between the atoms in all of the substances A to D. [4 marks]

d) Name the type of structure for each substance A to D. [4 marks]

6 The photographs show two uses of graphite. It is used in pencils and for electrodes in the electrolysis of sodium chloride solution.

![Figure 10.34]

a) With reference to the structure of graphite, explain why it is used in pencils. [3 marks]

b) Graphite electrodes conduct electricity. Explain why graphite is a good conductor of electricity. [2 marks]

c) Describe, as fully as you can, what happens when sodium atoms react with chlorine atoms to produce sodium chloride. You may use a diagram in your answer. [3 marks]

d) Explain why sodium chloride solution will conduct electricity, but sodium chloride solid will not. [2 marks]
7 The table gives some of the properties of the Period 3 element magnesium and one of its compounds, magnesium chloride.

Table 10.21

<table>
<thead>
<tr>
<th>Property</th>
<th>Magnesium</th>
<th>Magnesium chloride</th>
</tr>
</thead>
<tbody>
<tr>
<td>Melting point in °C</td>
<td>649</td>
<td>714</td>
</tr>
<tr>
<td>Electrical conductivity when solid</td>
<td>Conducts</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>Electrical conductivity when molten</td>
<td>Conducts</td>
<td>Conducts</td>
</tr>
</tbody>
</table>

Use ideas about structure and bonding to explain the similarities and differences between the properties of magnesium and magnesium chloride. [6 marks]

8 a) Chlorine is a green gas which exists as diatomic molecules.
   i) Suggest what is meant by the term diatomic. [1 mark]
   ii) Use a dot and cross diagram to clearly show how atoms of chlorine combine to form chlorine molecules. [2 marks]

b) Chlorine can form a range of compounds with both metals and non-metals.
   i) Describe, as fully as you can, what happens when calcium atoms react with chlorine atoms to produce calcium chloride. You may use a diagram in your answer. [3 marks]
   ii) Name the type of bonding found in calcium chloride. [1 mark]
   iii) Name the type of structure for calcium chloride. [1 mark]
   iv) Use a dot and cross diagram to show how atoms of chlorine combine with atoms of carbon to form tetrachloromethane CCl₄. [2 marks]
   v) Name the type of bonding found in CCl₄. [1 mark]

c) The properties of compounds depend very closely on their bonding. Redraw the following table with only the correct words to show some of the properties of calcium chloride and tetrachloromethane. [2 marks]

Table 10.22

<table>
<thead>
<tr>
<th>Compound</th>
<th>Solubility in water</th>
<th>Relative melting point</th>
</tr>
</thead>
<tbody>
<tr>
<td>Calcium chloride</td>
<td>Soluble/insoluble</td>
<td>Low/high</td>
</tr>
<tr>
<td>Tetrachloromethane</td>
<td>Soluble/insoluble</td>
<td>Low/high</td>
</tr>
</tbody>
</table>

d) The bonding in the elements calcium and carbon is very different. Describe the bonding in calcium and in carbon (in the form of graphite). [6 marks]

e) Both calcium and graphite can conduct electricity. State two properties of calcium which are different from those of graphite. [2 marks]

f) Why might fullerenes be used in new drug delivery systems? Choose the correct statement A, B or C. [1 mark]

A They are made from carbon atoms.
B They are hollow.
C They are very strong.

g) How does the structure of nanotubes make them suitable as catalysts? Choose the correct statement A, B or C. [1 mark]

A They have a large surface area to volume ratio.
B They are made from reactive carbon atoms.
C They have strong covalent bonds.
Working scientifically:
Units: Using prefixes and powers of ten for orders of magnitude

Standard form
Standard form is used to express very large or very small numbers so that they are more easily understood and managed. It is easier to say that a speck of dust has a mass of $1.2 \times 10^{-6}$ grams than to say it has a mass of 0.000 001 2 grams. It uses powers of ten.

Standard form must always look like this:

\[ A \times 10^n \]

A must always be between 1 and 10, and \( n \) is the number of places the decimal point moves (for numbers less than 1, the value of \( n \) will be negative).

Example
Write 4,600,000 in standard form.

Answer
Write the non-zero digits with a decimal place after the first number and then write \( \times 10 \) after it:

\[ 4.6 \times 10^6 \]

Then count how many places the decimal point has moved and write this as the \( n \) value. The \( n \) value is positive as 4,600,000 is greater than 1.

4,600,000 = $4.6 \times 10^6$

Example
Write 0.000,345 in standard form.

Answer
Write the non-zero digits with a decimal place after the first number and then write \( \times 10 \) after it:

\[ 3.45 \times 10^{-4} \]

Then count how many places the decimal point has moved and write this as the \( n \) value. The \( n \) value is negative as 0.000,345 is less than 1.

0.000,345 = $3.45 \times 10^{-4}$
Working scientifically: Units: Using prefixes and powers of ten for orders of magnitude

Figure 10.36 Make sure that you are familiar with how standard form is presented on your calculator. This calculator reads $1.23 \times 10^{99}$ to 3 significant figures.

SI units

The International System of Units (SI) is a system of units of measurements that is widely used all over the world, and the one which you will use in your study of chemistry. It uses several base units of measure, for example metres, grams and seconds. When a numerical unit is very small or large, the units may be modified by using a prefix. Some prefixes are shown in the table.

Table 10.23

<table>
<thead>
<tr>
<th>Prefix name</th>
<th>Prefix symbol</th>
<th>Scientific notation $10^n$</th>
<th>Decimal</th>
</tr>
</thead>
<tbody>
<tr>
<td>tera</td>
<td>T</td>
<td>$10^{12}$</td>
<td>1 000 000 000 000</td>
</tr>
<tr>
<td>giga</td>
<td>G</td>
<td>$10^9$</td>
<td>1 000 000 000</td>
</tr>
<tr>
<td>mega</td>
<td>M</td>
<td>$10^6$</td>
<td>1 000 000</td>
</tr>
<tr>
<td>kilo</td>
<td>k</td>
<td>$10^3$</td>
<td>1 000</td>
</tr>
<tr>
<td>centi</td>
<td>c</td>
<td>$10^{-2}$</td>
<td>0.01</td>
</tr>
<tr>
<td>milli</td>
<td>m</td>
<td>$10^{-3}$</td>
<td>0.001</td>
</tr>
<tr>
<td>micro</td>
<td>μ</td>
<td>$10^{-6}$</td>
<td>0.000 001</td>
</tr>
</tbody>
</table>

A prefix goes in front of a basic unit of measure to indicate a multiple of the unit. For example instead of writing 1000 grams we can add the prefix kilo ($10^3$) and write 1 kg.

For example using scientific notation:

- $1\text{ cm} = 1\text{ centimetre} = 1 \times 10^{-2}\text{ m} = 0.01\text{ m}$
- $1\text{ µl} = 1\text{ microlitre} = 1 \times 10^{-6}\text{ l} = 0.000 001\text{ l}$
- $1\text{ µm} = 1\text{ micrometre} = 1 \times 10^{-6}\text{ m} = 0.000 001\text{ m}$

Questions

1. Write the numbers below in standard form.
   a) 0.00024
   b) 32300000000
   c) 0.02
   d) 0.000000007
   e) 24000

2. Write the numbers below as decimals.
   a) $2.3 \times 10^{-3}$
   b) $4.6 \times 10^{2}$
   c) $9.5 \times 10^{-5}$
   d) $5.34 \times 10^{4}$
   e) $3.3 \times 10^{3}$

3. Write the following quantities in units with the appropriate prefixes.
   a) 310000000 m
   b) 0.001 g
   c) 9.700 m
   d) 0.000000002 s
Water companies regularly analyse samples of water supplied to homes to check that it is safe to drink. Their analytical chemists have to be able to carry out very accurate experiments to analyse the water. They also need to be able to carry out calculations using their results to work out how much of each substance is in the water. In this chapter you will learn how to perform a range of calculations.

This chapter covers specification points 5.3.1.1 to 5.3.2.5 and is called Quantitative chemistry.

It covers relative mass and moles, conservation of mass, and reacting. Related work on writing and balancing equations can be found in Chapter 14.
Relative mass and moles

Previously you could have learnt:

› Atoms are tiny.
› Nearly all of the mass of an atom is from the mass of the protons and neutrons.
› The mass number of an atom is the number of protons plus the number of neutrons in an atom.
› Mass is conserved in chemical reactions.

Test yourself on prior knowledge

1 What does the following equation tell you about the reaction:
   \[ \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \]

2 What is the principle of conservation of mass?

3 In a reaction, hydrogen reacts with oxygen to make water and nothing else. What mass of water will be made if 4 g of hydrogen reacts with 32 g of oxygen?

4 How many protons, neutrons and electrons are there in an atom of \(^{39}\text{K}\)?

Relative mass and moles

Relative atomic mass

Individual atoms have a tiny mass. For example, an atom of \(^{12}\text{C}\) has a mass of about \(2 \times 10^{-23}\) g (that is 0.000 000 000 000 000 000 000 02 g). As numbers like this are awkward to use, scientists measure the mass of atoms relative to each other. They use a scale where the mass of a \(^{12}\text{C}\) atom is defined as being exactly 12. On this scale, an atom of \(^{24}\text{Mg}\) has a relative mass of 24 and is twice as heavy as a \(^{12}\text{C}\) atom, whereas an atom of \(^{1}\text{H}\) has a relative mass of 1 and is 12 times lighter than a \(^{12}\text{C}\) atom. In effect, the relative mass of a single atom equals the mass number of that atom (mass number is the number of protons plus the number of neutrons).

Many elements are made up of a mixture of atoms of different isotopes. For example, 75% of chlorine atoms are \(^{35}\text{Cl}\) with relative mass of 35 and the remaining 25% are \(^{37}\text{Cl}\) atoms with relative mass of 37. The average relative mass of chlorine atoms is 35.5 as there are more chlorine atoms with relative mass of 35 than 37.

The relative atomic mass \((A_r)\) of an element is the average mass of atoms of that element taking into account the mass and amount of each isotope it contains.

Relative formula mass

The relative formula mass \((M_r)\) of a substance is the sum of the relative atomic masses of all the atoms shown in the formula (often referred to as formula mass).
For example, the formula of water is $\text{H}_2\text{O}$ and so the relative formula mass is the sum of the relative atomic mass of two hydrogen atoms ($2 \times 1$) and one oxygen atom (16) which adds up to 18. This and other examples are shown in Table 11.1.

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>$A_r$ values</th>
<th>Sum</th>
<th>$M_r$ (relative formula mass)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>$\text{H}_2\text{O}$</td>
<td>$H = 1, O = 16$</td>
<td>$2(1) + 16 = 18$</td>
<td>18</td>
</tr>
<tr>
<td>Copper</td>
<td>$\text{Cu}$</td>
<td>$\text{Cu} = 63.5$</td>
<td>$63.5$</td>
<td></td>
</tr>
<tr>
<td>Sodium chloride</td>
<td>$\text{NaCl}$</td>
<td>$\text{Na} = 23, \text{Cl} = 35.5$</td>
<td>$23 + 35.5 = 58.5$</td>
<td></td>
</tr>
<tr>
<td>Sulfuric acid</td>
<td>$\text{H}_2\text{SO}_4$</td>
<td>$H = 1, S = 32, O = 16$</td>
<td>$2(1) + 32 + 4(16) = 98$</td>
<td></td>
</tr>
<tr>
<td>Magnesium nitrate</td>
<td>$\text{Mg(NO}_3\text{)}_2$</td>
<td>$\text{Mg} = 24, \text{N} = 14, O = 16$</td>
<td>$24 + 2(14) + 6(16) = 148$</td>
<td></td>
</tr>
<tr>
<td>Ammonium sulfate</td>
<td>$(\text{NH}_4\text{)}_2\text{SO}_4$</td>
<td>$\text{N} = 14, H = 1, S = 32, O = 16$</td>
<td>$2(14) + 8(1) + 32 + 4(16) = 132$</td>
<td></td>
</tr>
</tbody>
</table>

The mole

What is a mole?

Amounts of chemicals are often measured in **moles**. This makes it much easier to work out how much of a chemical is needed in a reaction.

One atom of $^{12}\text{C}$ has a relative mass of 12. The mass of $6020000000000000000000000000$ ($6.02 \times 10^{23}$) atoms of $^{12}\text{C}$ is exactly 12 g. The number $6.02 \times 10^{23}$ is a very special number and is known as the **Avogadro constant**.

When you have a pair of $^{12}\text{C}$ atoms, you have two of them. When you have a dozen $^{12}\text{C}$ atoms, you have 12 of them. When you have a mole of $^{12}\text{C}$ atoms, you have $6.02 \times 10^{23}$ of them.

One mole of a substance contains the same number of the stated particles (atoms, molecules or ions) as one mole of any other substance. For example, the number of carbon atoms in one mole of carbon atoms ($6.02 \times 10^{23}$ atoms) is the same number of particles as there are molecules in one mole of water molecules ($6.02 \times 10^{23}$ molecules).

The value of the Avogadro constant was chosen so that the mass of one mole of that substance is equal to the relative formula mass ($M_r$) in grams. Table 11.2 shows some examples.
Table 11.2 Relative formula mass and the mole.

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>Relative formula mass ($M_r$)</th>
<th>Mass of 1 mole of that substance</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>H$_2$O</td>
<td>18</td>
<td>18 g</td>
</tr>
<tr>
<td>Copper</td>
<td>Cu</td>
<td>63.5</td>
<td>63.5 g</td>
</tr>
<tr>
<td>Sodium chloride</td>
<td>NaCl</td>
<td>58.5</td>
<td>58.5 g</td>
</tr>
<tr>
<td>Sulfuric acid</td>
<td>H$_2$SO$_4$</td>
<td>98</td>
<td>98 g</td>
</tr>
<tr>
<td>Magnesium nitrate</td>
<td>Mg(NO$_3$)$_2$</td>
<td>148</td>
<td>148 g</td>
</tr>
<tr>
<td>Ammonium sulfate</td>
<td>(NH$_4$)$_2$SO$_4$</td>
<td>132</td>
<td>132 g</td>
</tr>
</tbody>
</table>

How many moles?

If one mole of $^{12}$C atoms has a mass of 12 g, then it follows that the mass of two moles of $^{12}$C atoms will be 24 g. There is a simple equation linking the mass of a substance to the number of moles (Figure 11.2).

When using this equation, the mass must be in grams. Table 11.3 shows conversion factors if the masses are not given in grams.

Table 11.3 Conversion factors to work out mass in grams.

<table>
<thead>
<tr>
<th>Conversion factor</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 tonne = 1000000 g</td>
<td>3 tonnes = 3 × 1000000 = 3 000 000 g</td>
</tr>
<tr>
<td>1 kg = 1000 g</td>
<td>0.5 kg = 0.5 × 1000 = 500 g</td>
</tr>
<tr>
<td>1 mg = 0.001 g</td>
<td>20 mg = 20 × 0.001 = 0.020 g</td>
</tr>
</tbody>
</table>

Table 11.4 gives some examples using the equation that links mass and moles. In calculations, the units of moles is usually abbreviated to mol.

Table 11.4 Examples of working out mass and number of moles.

<table>
<thead>
<tr>
<th>How many moles in each of the following?</th>
<th>180 g of H$_2$O</th>
<th>4 g of CH$_4$</th>
<th>2 kg of Fe$_2$O$_3$</th>
<th>50 mg of NaOH</th>
<th>20 moles of CO$_2$</th>
<th>0.025 moles of Cl$_2$</th>
<th>3.6 g of a substance is found to contain 0.020 mol</th>
</tr>
</thead>
<tbody>
<tr>
<td>$M_r$ of H$_2$O = 2(1) + 16 = 18</td>
<td>Moles H$_2$O = $\frac{mass}{M_r}$ = $\frac{180}{18}$ = 10 mol</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$M_r$ of CH$_4$ = 12 + 4(1) = 16</td>
<td>Moles CH$_4$ = $\frac{mass}{M_r}$ = $\frac{4}{16}$ = 0.25 mol</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$M_r$ of Fe$_2$O$_3$ = 2(56) + 3(16) = 160</td>
<td>Mass of Fe$_2$O$_3$ = 2 kg = 2 × 1000 g = 2000 g</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$M_r$ of NaOH = 23 + 16 + 1 = 40</td>
<td>mass NaOH = $\frac{50}{1000}$ = 0.050 g</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$M_r$ of CO$_2$ = 12 + 2(16) = 44</td>
<td>Moles CO$_2$ = $M_r$ × moles = 44 × 20 = 880 g</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$M_r$ of Cl$_2$ = 2(35.5) = 71</td>
<td>Moles Cl$_2$ = $M_r$ × moles = 71 × 0.025 = 1.78 g</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$M_r$ = $\frac{mass}{moles}$ = $\frac{3.6}{0.020}$ = 180</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Test yourself

You can find relative atomic masses on the periodic table in the Appendix to help you answer these questions.

2 What is the mass of one mole of the following substances?
   a) iron, Fe  
   b) oxygen, O₂  
   c) ethane, C₂H₆  
   d) potassium chloride, KCl  
   e) calcium nitrate, Ca(NO₃)₂

3 Which one of the substances in each of the following pairs of substances contains the most particles or are they the same?
   a) 2 moles of water (H₂O) molecules and 2 moles of carbon dioxide (CO₂) molecules
   b) 10 moles of methane (CH₄) molecules and 10 moles of argon atoms (Ar)
   c) 10 moles of helium atoms (He) and 5 moles of oxygen molecules (O₂)

4 What are the following masses in grams?
   a) 20 kg  
   b) 5 mg  
   c) 0.3 tonnes

5 What is the mass of each of the following?
   a) 3 moles of oxygen, O₂  
   b) 0.10 moles of ethanol, C₂H₅OH

6 How many moles are there in each of the following?
   a) 50 g of hydrogen, H₂  
   b) 4 kg of calcium carbonate, CaCO₃  
   c) 80 mg of bromine, Br₂  
   d) 2 tonnes of calcium oxide, CaO

7 0.300 g of a substance was analysed and found to contain 0.0050 moles. Calculate the $M_r$ of the substance.

8 Give the following numbers to 2, 3 and 4 significant figures.
   a) 34.822 6  
   b) 28554.210  
   c) 0.023 187 6  
   d) 0.000631947

9 Find the mean of these measurements and give your answer to 3 significant figures.
   a) 25.4 cm³, 25.1 cm³, 25.3 cm³  
   b) 162 s, 175 s, 169 s, 173 s  
   c) 1.65 g, 1.70 g, 1.69 g, 1.64 g, 1.71 g

Show you can...

Use Figure 11.2 to write three different equations linking mass, moles and $M_r$. Use these equations to answer the questions below.

a) Calculate the number of moles in 9.8 g of H₂SO₄.

b) Calculate the mass in grams of 0.5 moles of Ca(OH)₂.

c) 6.9 g of a substance Y₂CO₃ contains 0.05 moles. What is the relative formula mass of the substance? Identify Y.

Table 11.5 Some numbers to 2, 3 and 4 significant figures.

<table>
<thead>
<tr>
<th>Number</th>
<th>2 sf</th>
<th>3 sf</th>
<th>4 sf</th>
</tr>
</thead>
<tbody>
<tr>
<td>2.7358</td>
<td>2.7</td>
<td>2.74</td>
<td>2.736</td>
</tr>
<tr>
<td>604531</td>
<td>600000</td>
<td>605000</td>
<td>604500</td>
</tr>
<tr>
<td>0.10836</td>
<td>0.11</td>
<td>0.108</td>
<td>0.1084</td>
</tr>
<tr>
<td>0.0042981</td>
<td>0.0043</td>
<td>0.00430</td>
<td>0.004298</td>
</tr>
</tbody>
</table>

○ Significant figures

We often quote answers to calculations to a certain number of significant figures. In chemistry, we usually give values to 2, 3 or 4 significant figures (sf), but it can be more or less than this. Table 11.5 shows some numbers given to 2, 3 and 4 significant figures.

We quote values to a limited number of significant figures because we cannot be sure of the exact value to a greater number of significant figures.

For example, if we measure the temperature rise in a reaction three times and find the values to be 21°C, 21°C and 22°C, then the mean temperature rise shown on a calculator would be 21.333 333 33°C. However, it is impossible for us to say that the temperature rise is exactly 21.333 333 33°C as the thermometer could only measure to ±1°C. Therefore, we should quote the temperature rise to 2 significant figures, i.e. 21°C, which is the same number of significant figures as the values we measured.
Conservation of mass

Balanced equations
In a balanced chemical equation, the number of atoms of each element is the same on both sides of the equation. This is because atoms cannot be created or destroyed in chemical reactions. This means that you have the same atoms before and after the reaction, although how they are bonded to each other changes during the reaction.

For example, in Figure 11.3 one molecule of nitrogen (N₂) reacts with three molecules of hydrogen (H₂) to make two molecules of ammonia (NH₃).

There are two N atoms and six H atoms on both sides of the equation.

There is a section on how to balance symbol equations later in this chapter.

Conservation of mass
As there are the same atoms present at the start and end of a chemical reaction, mass must be conserved in the reaction. In other words, the total mass of the reactants must equal the total mass of the products. This is known as the law of conservation of mass.

One way to look at this is using relative formula masses. The total of the relative formula masses of all the reactants in the quantities shown in the equation will add up to the total of the relative formula masses of all the products in the quantities shown in the equation (Figure 11.4).

All chemical reactions obey the law of conservation of mass. There are some reactions that may appear to break this law, but they do not as shown below.

Reaction of metals with oxygen
When metals react with oxygen the mass of the product is greater than the mass of the original metal. However, this does not break the law of conservation of mass. The ‘extra’ mass is the mass of the oxygen from the air that has combined with the metal to form a metal oxide.

In the example shown in Figure 11.5, some magnesium has been heated in a crucible. It may appear as though the products are 0.16 g heavier than the reactants, but 0.16 g of oxygen has reacted with the 0.24 g magnesium to make 0.40 g magnesium oxide.
**Thermal decomposition reactions**

A thermal decomposition reaction is one where high temperature causes a substance to break down into simpler substances. In decomposition reactions, one or more of the products may escape from the reaction container into the air as a gas. These reactions may appear to lose mass, but the ‘missing’ mass is the mass of the gas that has escaped into the air.

In the example shown in Figure 11.6, some copper carbonate is heated in a crucible. It may appear as though the products are 0.22 g lighter than the reactants, but 0.22 g of carbon dioxide has been released into the air.

**KEY TERM**

**Thermal decomposition** Reaction where high temperature causes a substance to break down into simpler substances.

**TIP**

Metal carbonates decompose into a metal oxide and carbon dioxide when heated.

---

**Oxidation of titanium**

The metal titanium reacts with oxygen to form an oxide of titanium. In an experiment a sample of titanium metal was heated in a crucible with a lid. During heating the lid was lifted from time to time.

The following results were obtained:
- Mass of crucible = 16.34 g
- Mass of crucible + titanium metal = 17.36 g
- Mass of crucible + titanium oxide = 18.04 g

**Questions**
1. Use the results to calculate
   a) the mass of titanium used in this experiment
   b) the mass of titanium oxide formed in this experiment
   c) the mass of oxygen used in this experiment.
2. The equation for the reaction is Ti + O₂ → TiO₂
   Explain using the masses calculated in (1) how this reaction follows the law of conservation of mass.
3. Suggest why it was necessary to lift the crucible lid during heating.
   In a different experiment titanium metal was heated in a stream of oxygen:

**Questions**
4. What masses should be found before heating to determine the mass of titanium used in the experiment?
5. The ceramic dish and its contents are repeatedly weighed, heated, reweighed and heated until the mass is constant. State and explain if the mass increases or decreases during this experiment.
6. What safety precautions should be taken in this experiment?
7. How would the repeatability of this experiment be checked?
Balancing symbol equations

To balance symbol equations, you need to be able to work out:

- the number of elements in a chemical formula; for example, NaCl has two elements – Na (sodium) and Cl (chlorine); similarly, NaOH has three different elements – Na (sodium), O (oxygen) and H (hydrogen)
- the number of atoms; NaCl has one atom of sodium (Na), and one atom of chlorine (Cl); however, water (H₂O) has two atoms of hydrogen and one of oxygen.

When working out the number of atoms you should remember that if the number is subscript (the small number at the bottom), it refers to the element immediately before it; for example, in 2NaHCO₃ the subscript 3 means there are three atoms of oxygen (O).

If the number is normal size, for example 2NaHCO₃, it refers to all the elements after it. In this example there are two sodium (Na) atoms, two hydrogen (H), two carbon (C) and six oxygen (O) atoms (2 × 3).

The formula Ca(HCO₃)₂ has one atom of calcium, two of hydrogen, two of carbon and six of oxygen (the brackets indicates that everything inside the bracket is multiplied by the number in subscript outside).

Example

The symbol equation for the reaction between sodium hydroxide and hydrochloric acid is

NaOH + HCl → NaCl + H₂O

This equation is balanced as there are the same numbers of each type of atom on either side.

Example

Magnesium burns in oxygen to give magnesium oxide.

Mg + O₂ → MgO

In this equation there are two oxygen atoms on the left side and only one on the right.

We need to double the number of oxygens on the right by using the balancing number 2, i.e. 2MgO, but as we have also doubled the number of magnesium atoms we need to balance them too. We do this by doubling the Mg on the left.

2Mg + O₂ → 2MgO
Example
When heated sodium hydrogencarbonate breaks down to form sodium carbonate, carbon dioxide and water.

The unbalanced symbol equation is:

\[ \text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O} \]

The balanced symbol equation for heating sodium hydrogencarbonate is:

\[ 2\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O} \]

We know it is balanced because:
- there are two Na atoms on the left (2Na) and two on the right (Na₂)
- two H atoms on the left and two on the right
- two C atoms on the left and two on the right
- six O atoms on the left and six on the right.

Example
Lithium reacts with water to give lithium hydroxide and hydrogen.

Using symbols the unbalanced equation is:

\[ 3\text{Li} + \text{H}_2\text{O} \rightarrow \text{LiOH} + \text{H}_2 \]

This equation does need to be balanced, as there are different numbers of atoms of hydrogen on each side.

To balance the number of atoms on either side, we need to balance the hydrogen atoms (currently two on the left and three on the right). We do this by using balancing numbers as shown below to give the balanced symbol equation:

\[ 2\text{Li} + 2\text{H}_2\text{O} \rightarrow 2\text{LiOH} + \text{H}_2 \]

State symbols
Balanced equations sometimes state symbols to show the state of each substance: (s) solid, (l) liquid, (g) gas, (aq) aqueous. For example:

\[ \text{CaCO}_3(s) + 2\text{HCl}(aq) \rightarrow \text{CaCl}_2(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g) \]
Test yourself

You can find relative atomic masses on the periodic table in the Appendix to help you answer these questions.

10 Hydrogen reacts with oxygen to make water as shown in this equation below.
\[ 2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \]

a) Describe in words what this tells you about the reaction of hydrogen with oxygen in terms of how many molecules are involved in the reaction.

b) Show that the sum of the relative formula masses of all the reactants equals the sum of the relative formula masses of all the products.

11 A piece of copper was heated in air. After a few minutes it was reweighed and found to be heavier.

a) Explain why the copper gets heavier.

b) What is the law of conservation of mass?

c) Explain why this reaction does not break the law of conservation of mass.

12 When 1.19 g of solid nickel carbonate is heated for several minutes, only 0.75 g of solid remains. Explain clearly why the mass decreases and what happens to the remaining mass.

Reacting masses

Molar ratios in equations

Chemical equations can be interpreted in terms of moles. For example, in the equation in Figure 11.9 one mole of nitrogen (N\text{2}) molecules reacts with three moles of hydrogen (H\text{2}) molecules to form two moles of ammonia (NH\text{3}) molecules.

\[ \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \]

TIP
The Avogadro constant is the number of particles in one mole of a substance and has a value of \(6.02 \times 10^{23}\).
These ratios can be used to calculate how many moles would react and be produced in reactions (Figures 11.10 and 11.11).

\[
C_3H_8 + 3O_2 \rightarrow 2CO_2 + 4H_2O
\]

\[
C_2H_4 + 2O_2 \rightarrow CO_2 + 2H_2O
\]

\[
Be + Cl_2 \rightarrow BeCl_2
\]

\[
C_2H_6 + 3O_2 \rightarrow 2CO_2 + 2H_2O
\]

\[
2KClO_3 \rightarrow 2KCl + 3O_2
\]

Test yourself

13 Hydrogen reacts with oxygen to make water as shown in this equation below.
\[2H_2 + O_2 \rightarrow 2H_2O\]

a) How many moles of oxygen react with 10 moles of hydrogen?
b) How many moles of hydrogen react with 3 moles of oxygen?
c) How many moles of oxygen react with 0.3 moles of hydrogen?
d) How many moles of hydrogen react with 2.5 moles of oxygen?

14 Titanium is made when titanium chloride reacts with sodium as shown in the equation below.
\[TiCl_4 + 4Na \rightarrow Ti + 4NaCl\]

a) How many moles of sodium react with 4 moles of titanium chloride?
b) How many moles of titanium are made from 2.5 moles of titanium chloride?
c) How many moles of sodium react with 0.5 moles of titanium chloride?
d) How many moles of titanium are made from 0.65 moles of titanium chloride?

15 Potassium chlorate \((KClO_3)\) decomposes to form potassium chloride and oxygen as shown below.
\[2KClO_3 \rightarrow 2KCl + 3O_2\]

a) How many moles of potassium chlorate are formed when 10 moles of potassium chlorate decomposes?
b) How many moles of oxygen are formed when 4 moles of potassium chlorate decomposes?
c) How many moles of potassium chlorate are formed when 0.5 moles of potassium chlorate decomposes?
d) How many moles of oxygen are formed when 3 moles of potassium chlorate decomposes?

Calculating reacting masses

Scientists need to be able to calculate how much of each substance to use in a chemical reaction. There are two common ways to do these calculations. One method uses moles while the other uses ratios and relative formula masses.
Reacting masses

Examples

Using moles
1 Work out the relative formula mass ($M_r$) of the substance whose mass is given and the one you are finding the mass of. (Remember that the balancing numbers in the equation are not part of the formulae).
2 Calculate the moles of the substance whose mass is given (using moles = mass / $M_r$).
3 Use molar ratios from the balanced equation to work out the moles of the substance the question asks about.
4 Calculate the mass of that substance (using mass = $M_r \times$ moles).

Example 1
Calculate the mass of hydrogen needed to react with 140 g of nitrogen to make ammonia.

\[ \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \]

$M_r$ N$_2$ = 2(14) = 28, $M_r$ H$_2$ = 2(1) = 2

Moles N$_2$ = \( \frac{\text{mass}}{M_r} = \frac{140}{28} = 5 \)

Moles H$_2$ = moles of N$_2$ \times 3 = 5 \times 3 = 15

Mass H$_2$ = $M_r$ \times moles = 2 \times 15 = 30 g

Using ratios
1 Work out the relative formula mass ($M_r$) of the substance whose mass is given and the one you are finding the mass of. (Remember that the balancing numbers in the equation are not part of the formulae).
2 Find the reacting mass ratio for these substances using the $M_r$ values and the balancing numbers in the equation.
3 Scale this to find what happens with 1 g of the substance given.
4 Scale this up/down to the mass you were actually given.

Example 1
Calculate the mass of hydrogen needed to react with 140 g of nitrogen to make ammonia.

$N_2 + 3H_2 \rightarrow 2NH_3$

$M_r$ N$_2$ = 2(14) = 28, $M_r$ H$_2$ = 2(1) = 2

N$_2$ reacts with 3H$_2$

28 g of N$_2$ reacts with 6 g (3 \times 2) of H$_2$

1 g of N$_2$ reacts with \( \frac{6}{28} \) g of H$_2$

140 g of N$_2$ reacts with 140 \times \( \frac{6}{28} \) = 30 g of H$_2$

Example 2
What mass of oxygen reacts with 4.6 g of sodium?

$4\text{Na} + \text{O}_2 \rightarrow 2\text{Na}_2\text{O}$

$M_r$ Na = 23, $M_r$ O$_2$ = 2(16) = 32

Moles Na = \( \frac{\text{mass}}{M_r} = \frac{4.6}{23} = 0.20 \)

Moles O$_2$ = moles of Na ÷ 4 = 0.20 ÷ 4 = 0.05

Mass O$_2$ = $M_r$ \times moles = 32 \times 0.05 = 1.6 g

Example 3
What mass of iron is produced when 32 kg of iron oxide is heated with carbon monoxide?

$\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$

$M_r$ Fe$_2$O$_3$ = 2(56) + 3(16) = 160, $M_r$ Fe = 56

Moles Fe$_2$O$_3$ = \( \frac{\text{mass}}{M_r} = \frac{32000}{160} = 200 \)

Moles Fe = moles of Fe$_2$O$_3$ \times 2 = 200 \times 2 = 400

Mass Fe = $M_r$ \times moles = 56 \times 400 = 22400 g

Test yourself

16 Hydrogen reacts with oxygen to make water as shown in the equation below. What mass of oxygen is needed to react with 10 g of hydrogen?

$2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$

17 Calcium carbonate decomposes when heated as shown in the equation below. What mass of calcium oxide is formed from 25 g of calcium carbonate?

$\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$

18 Magnesium burns in oxygen to form magnesium oxide as shown in the equation below. What mass of magnesium oxide would be made from 3 g of magnesium?

$2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$

19 Lithium reacts with water as shown in the equation below. What mass of hydrogen would be made from 1.4 g of lithium?

$2\text{Li} + 2\text{H}_2\text{O} \rightarrow 2\text{LiOH} + \text{H}_2$

20 Aluminium is made by the electrolysis of aluminium oxide as shown in the equation below. What mass of aluminium can be formed from 1 kg of aluminium oxide? Give your answer to 3 significant figures.

$2\text{Al}_2\text{O}_3 \rightarrow 4\text{Al} + 3\text{O}_2$

21 Hot molten copper can be produced for welding electrical connections in circuits by the reaction of copper oxide with aluminium powder shown below. What mass of aluminium is needed to react with 10 g of copper oxide? Give your answer to 3 significant figures.

$3\text{CuO} + 2\text{Al} \rightarrow 3\text{Cu} + \text{Al}_2\text{O}_3$
Deducing the balancing numbers in an equation from reacting masses

The balancing numbers in a chemical equation can be calculated by calculating the moles of the substances in the reaction. In order to do this:

1. Calculate the moles of each substance (using moles = \( \frac{\text{Mass}}{M_r} \)).
2. Find the simplest whole number ratio of these mole values by dividing all the mole values by the smallest mole value.
3. If this does not give a whole number ratio, multiply up by a factor of 2 (where there is a value ending in approximately 0.5), of 3 (where there is a value ending in approximately 0.33 or 0.67), of 4 (where there is a value ending in approximately 0.25 or 0.75), etc.

Example

1.2 g of magnesium (Mg) reacts with 0.8 g of oxygen (O\(_2\)) to make 2.0 g of magnesium oxide (MgO). Use this information to deduce the equation for this reaction.

**Answer**

<table>
<thead>
<tr>
<th>Substance</th>
<th>Magnesium (Mg)</th>
<th>Oxygen (O(_2))</th>
<th>Magnesium oxide (MgO)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Calculate the moles of each substance</td>
<td>moles = ( \frac{\text{mass}}{M_r} )</td>
<td>moles = ( \frac{\text{mass}}{M_r} )</td>
<td>moles = ( \frac{\text{mass}}{M_r} )</td>
</tr>
<tr>
<td></td>
<td>( \frac{1.2}{24} = 0.05 )</td>
<td>( \frac{0.8}{32} = 0.025 )</td>
<td>( \frac{2.0}{40} = 0.05 )</td>
</tr>
<tr>
<td>Find the simplest whole number ratio</td>
<td>( \frac{0.05}{0.025} = 2 )</td>
<td>( \frac{0.025}{0.025} = 1 )</td>
<td>( \frac{0.05}{0.025} = 2 )</td>
</tr>
</tbody>
</table>

Therefore the reacting ratio is 2:1:2, and so the balanced equation is

\[ 2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO} \]

Example

0.81 g of aluminium (Al) reacts with 3.20 g of chlorine (Cl\(_2\)) to make 4.01 g of aluminium chloride (AlCl\(_3\)). Use this information to deduce the equation for this reaction.

**Answer**

<table>
<thead>
<tr>
<th>Substance</th>
<th>Aluminium (Al)</th>
<th>Chlorine (Cl(_2))</th>
<th>Aluminium chloride (AlCl(_3))</th>
</tr>
</thead>
<tbody>
<tr>
<td>Calculate the moles of each substance</td>
<td>moles = ( \frac{\text{mass}}{M_r} )</td>
<td>moles = ( \frac{\text{mass}}{M_r} )</td>
<td>moles = ( \frac{\text{mass}}{M_r} )</td>
</tr>
<tr>
<td></td>
<td>( \frac{0.81}{27} = 0.030 )</td>
<td>( \frac{3.2}{71} = 0.045 )</td>
<td>( \frac{4.01}{133.5} = 0.030 )</td>
</tr>
<tr>
<td>Find the simplest whole number ratio</td>
<td>( \frac{0.030}{0.030} \times 2 = 2 )</td>
<td>( \frac{0.045}{0.030} \times 2 = 3 )</td>
<td>( \frac{0.030}{0.030} \times 2 = 2 )</td>
</tr>
</tbody>
</table>

Therefore the reacting ratio is 2:3:2, and so the balanced equation is

\[ 2\text{Al} + 3\text{Cl}_2 \rightarrow 2\text{AlCl}_3 \]
Example

2.2 g of propane (C₃H₈) reacts with 8.0 g of oxygen (O₂). Calculate the molar ratio in which propane and oxygen react here.

<table>
<thead>
<tr>
<th>Substance</th>
<th>propane (C₃H₈)</th>
<th>oxygen (O₂)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Calculate the moles of each substance</td>
<td>moles = mass/M = 2.2/44 = 0.050</td>
<td>moles = mass/M = 8.0/32 = 0.25</td>
</tr>
<tr>
<td>Find the simplest whole number ratio</td>
<td>0.050/0.050 = 1</td>
<td>0.25/0.050 = 5</td>
</tr>
</tbody>
</table>

Therefore the reacting ratio is 1:5, and so the ratio they react in is C₃H₈ + 5O₂.

Test yourself

22 6.8 g of hydrogen peroxide (H₂O₂) decomposes into 3.6 g of water (H₂O) and 3.2 g of oxygen (O₂). By calculating molar ratios, deduce the balanced equation for this reaction.

23 2.0 g of calcium (Ca) reacts with 1.9 g of fluorine (F₂) to form 3.9 g of calcium fluoride (CaF₂). By calculating molar ratios, deduce the balanced equation for this reaction.

24 1.0 g of nickel (Ni) reacts with 0.27 g of oxygen (O₂) to form 1.27 g of nickel oxide (NiO). By calculating molar ratios, deduce the balanced equation for this reaction.

25 When 4.1 g of calcium nitrate (Ca(NO₃)₂) is heated, 1.4 g of calcium oxide (CaO), 2.3 g of nitrogen dioxide (NO₂) and 0.4 g of oxygen (O₂) are formed. By calculating molar ratios, deduce the balanced equation for this reaction.

26 11.7 g of potassium (K) reacts with 24 g of bromine (Br₂). Calculate the molar ratio in which potassium reacts with bromine.

27 11.4 g of titanium chloride (TiCl₄) reacts with 5.52 g of sodium (Na). Calculate the molar ratio in which titanium chloride reacts with sodium.

Experiment to find the equation for the action of heating sodium hydrogencarbonate

A student suggested that there are three possible equations for the thermal decomposition of sodium hydrogencarbonate.

Equation 1: NaHCO₃ → NaOH + CO₂
Equation 2: 2NaHCO₃ → Na₂CO₃ + H₂O + CO₂
Equation 3: 2NaHCO₃ → Na₂O + H₂O + 2CO₂

In order to find out which is correct, she carried out the following experiment and recorded her results.

- Find the mass of an empty evaporating basin.
- Add approximately 8 g of sodium hydrogencarbonate to the basin and find the mass.
- Heat gently for about 5 minutes.
- Allow to cool and then find the mass.
- Reheat, cool and find the mass.
- Repeat the heating and measurement of mass until constant mass is obtained.

Results

| Mass of evaporating basin = 21.05 g |
| Mass of basin and sodium hydrogencarbonate = 29.06 g |
| Mass of sodium hydrogencarbonate = 8.01 g |
| Mass of basin and residue after heating to constant mass = 26.10 g |
| Mass of residue = 5.05 g |

Questions

1. Draw a labelled diagram of the assembled apparatus for this experiment.
2. Calculate the number of moles of sodium hydrogencarbonate used.
3. From the possible equations, and your answer to question 2, calculate:
   a) The number of moles of NaOH that would be formed in equation 1.
   b) The number of moles of Na₂CO₃ that would be formed in equation 2.
   c) The number of moles of Na₂O that would be formed in equation 3.
4. From your answers to question 3, calculate:
   a) The mass of NaOH that would be formed in equation 1.
   b) The mass of Na₂CO₃ that would be formed in equation 2.
   c) The mass of Na₂O that would be formed in equation 3.
5. By comparing your answers in question 4 with the experimental mass of residue, deduce which is the correct equation for the decomposition of sodium hydrogencarbonate.
6. Why did the mass decrease?
● Using an excess

In many reactions involving two reactants, it is very common for an excess of one of the reactants to be used to ensure that all of the other reactant is used up. This is often done if one of the reactants is readily available but the other one is expensive or is in limited supply. For example, when many fuels are burned an excess of oxygen is used. Fuels are expensive and in limited supply. The oxygen is readily available from the air and using an excess of oxygen ensures that all the fuel burns.

When one of the reactants is in excess, the other reactant is a limiting reactant that is completely used up. This is because it is the amount of this substance that determines the amount of product formed in a reaction, in other words it limits the amount of product made.

**Example**

Magnesium reacts with sulfuric acid as shown below: 5 moles of magnesium (Mg) is reacted with 7 moles of sulfuric acid (H₂SO₄). One of the reagents is in excess. Calculate the moles of the products formed.

\[
\text{Mg} + \text{H}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{H}_2
\]

**Answer**

As magnesium and sulfuric acid react with the ratio 1 to 1, only 5 moles of the sulfuric acid can react as there are only 5 moles of magnesium to react with. The rest of the sulfuric acid is in excess (2 moles). Therefore, the magnesium is the limiting reactant and determines how much of the products are made. In this case, 5 moles of MgSO₄ and 5 moles of H₂ would be made.

<table>
<thead>
<tr>
<th>Table 11.7</th>
</tr>
</thead>
<tbody>
<tr>
<td>Reacting ratio from the equation</td>
</tr>
<tr>
<td>Amount provided</td>
</tr>
<tr>
<td>Reaction that takes place</td>
</tr>
</tbody>
</table>

**Example**

Iron oxide reacts with carbon monoxide as shown below to produce iron. 10 moles of iron oxide (Fe₂O₃) is reacted with 50 moles of carbon monoxide (CO). Calculate the moles of products formed.

\[
\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2
\]

**Answer**

In this reaction, the reacting ratio is one mole of iron oxide for every three moles of carbon monoxide. Therefore, 10 moles of iron oxide reacts with 30 moles of carbon monoxide. The rest of the carbon monoxide is in excess (20 moles). Therefore, the iron oxide is the limiting reactant and determines how much of the products are made. In this case, 20 moles of iron and 30 moles of carbon dioxide would be made.

<table>
<thead>
<tr>
<th>Table 11.8</th>
</tr>
</thead>
<tbody>
<tr>
<td>Reacting ratio from the equation</td>
</tr>
<tr>
<td>Amount provided</td>
</tr>
<tr>
<td>Reaction that takes place</td>
</tr>
</tbody>
</table>

**KEY TERMS**

- **Excess** When the amount of a reactant is greater than the amount that can react.
- **Limiting reactant** The reactant in a reaction that determines the amount of products formed. Any other reagents are in excess and will not all react.
Example
Magnesium reacts with oxygen as shown below. 0.30 moles of magnesium (Mg) is reacted with 0.20 moles of oxygen (O₂). Calculate the moles of products formed.

\[ 2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO} \]

**Answer**
In this reaction, the reacting ratio is two moles of magnesium for every mole of oxygen. Therefore, 0.30 moles of magnesium can only react with 0.15 moles of oxygen. The rest of the oxygen is in excess (0.05 moles). Therefore, the magnesium is the limiting reactant and determines how much of the products is made. In this case, 0.30 moles of MgO would be made.

<table>
<thead>
<tr>
<th>Reacting ratio from the equation</th>
<th>2 moles of Mg</th>
<th>reacts with 1 mole of O₂</th>
<th>to make 2 moles of MgO</th>
</tr>
</thead>
<tbody>
<tr>
<td>Amount provided</td>
<td>0.30 moles of Mg</td>
<td>0.20 moles of O₂</td>
<td>limiting reactant</td>
</tr>
<tr>
<td>Reaction that takes place</td>
<td>0.30 moles of Mg</td>
<td>reacts with 0.15 moles of O₂</td>
<td>to make 0.30 moles of MgO</td>
</tr>
</tbody>
</table>

Example
Tungsten oxide reacts with hydrogen as shown below to produce tungsten (Figure 11.13). 23.2 g of tungsten oxide (WO₃, \( M_r = 232 \)) is reacted with 20.0 g of hydrogen (H₂, \( M_r = 2 \)). Calculate the mass of tungsten (\( M_r = 184 \)) formed.

\[ \text{WO}_3 + 3\text{H}_2 \rightarrow \text{W} + 3\text{H}_2\text{O} \]

**Answer**
In this reaction, the reacting ratio is one mole of tungsten oxide for every three moles of hydrogen. Therefore, 0.10 moles (calculated from 23.2 g) of tungsten oxide reacts with 0.30 moles of hydrogen, and so the hydrogen is in excess (there are 10 moles of hydrogen calculated from 20 g). Therefore, the tungsten oxide is the limiting reactant and determines how much of the products are made. In this case, 0.10 moles of tungsten and 0.30 moles of water would be made.

<table>
<thead>
<tr>
<th>Reacting ratio from the equation</th>
<th>1 mole of WO₃</th>
<th>reacts with 3 moles of H₂</th>
<th>to make 1 mole of W and 3 moles of H₂O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Amount provided</td>
<td>Moles = ( \frac{23.2}{232} = 0.10 )</td>
<td>Moles = ( \frac{20}{10} = 10 )</td>
<td>limiting reactant</td>
</tr>
<tr>
<td>Reaction that takes place</td>
<td>0.10 moles of WO₃</td>
<td>reacts with 0.30 moles of H₂</td>
<td>to make 0.10 moles of W and 0.30 moles of H₂O</td>
</tr>
</tbody>
</table>
Test yourself

28 Copper can be made by reacting copper oxide with hydrogen as shown below.
\[ \text{CuO} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{O} \]
How many moles of copper would be made if:
- a) 5 moles of copper oxide were reacted with 10 moles of hydrogen?
- b) 2 moles of copper oxide were reacted with 2 moles of hydrogen?
- c) 0.4 moles of copper oxide were reacted with 0.3 moles of hydrogen?

29 Copper can be also be made by reacting copper oxide with methane as shown below.
\[ 4\text{CuO} + \text{CH}_4 \rightarrow 4\text{Cu} + 2\text{H}_2\text{O} + \text{CO}_2 \]
How many moles of copper would be made if:
- a) 2 moles of copper oxide were reacted with 2 moles of methane?
- b) 2 moles of copper oxide were reacted with 1 mole of methane?
- c) 10 moles of copper oxide were reacted with 2 moles of methane?

30 Hydrogen gas is formed when magnesium reacts with hydrochloric acid as shown below.
\[ \text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2 \]
How many moles of hydrogen would be made if:
- a) 3 moles of magnesium were reacted with 3 moles of hydrochloric acid?
- b) 0.2 moles of magnesium were reacted with 0.3 moles of hydrochloric acid?
- c) 0.5 moles of magnesium were reacted with 1.5 moles of hydrochloric acid?

31 Calcium reacts with sulfur as shown in the equation below. What mass of calcium sulfide can be made when 8 g of calcium reacts with 8 g of sulfur?
\[ \text{Ca} + \text{S} \rightarrow \text{CaS} \]

32 Titanium is made when titanium chloride reacts with magnesium as shown in the equation below. What mass of titanium can be made when 1.9 g of titanium chloride reacts with 6 g of magnesium?
\[ \text{TiCl}_4 + 2\text{Mg} \rightarrow \text{Ti} + 2\text{MgCl}_2 \]

33 Aluminium bromide is made when aluminium reacts with bromine as shown in the equation below. What mass of aluminium bromide can be made when 0.81 g of aluminium reacts with 6.4 g of bromine?
\[ 2\text{Al} + 3\text{Br}_2 \rightarrow 2\text{AlBr}_3 \]

Show you can...

Calculate the mass of calcium carbide (CaC\(_2\)) formed when 84 g of calcium oxide (CaO) is reacted with 48 g of coke (C).
\[ \text{CaO} + 3\text{C} \rightarrow \text{CaC}_2 + \text{CO} \]

Use the following headings.
- Number of moles of calcium oxide
- Number of moles of coke
- The reactant in excess is
- Number of moles of calcium carbide formed
- Mass of calcium carbide formed in grams.
The concentration of solutions

Concentration of solutions in g/dm³

Figure 11.14 shows two solutions of copper sulfate. The one that is darker blue has much more copper sulfate dissolved in it. The darker blue one is more concentrated and the paler blue one is more dilute.

We can measure the concentration of a solution by considering what mass of solute is dissolved in the solution. This is usually found in g/dm³, which means the number of grams of solute dissolved in each dm³ of solution (Figure 11.15); 1 dm³ is the same volume as 1000 cm³ or 1 litre. For example, if 50 grams of copper sulfate is dissolved in 2 dm³ of solution, then the concentration is 25 g/dm³.

\[
\text{concentration (g/dm}^3\text{)} = \frac{\text{mass dissolved (g)}}{\text{volume (dm}^3\text{)}}
\]

This formula triangle may help you. Cover up the quantity you want, to show the equation you need to use.

In the laboratory, we often use volumes measured in cm³ rather than dm³. As there are 1000 cm³ in 1 dm³, we should divide the volume in cm³ by 1000 to find the volume in dm³. For example, 25 cm³ is \(\frac{25}{1000} = 0.025\) dm³.

TIP

1 dm³ = 1 litre = 1000 cm³.
Chapter review questions

1. Calculate the relative formula mass ($M_r$) of the following substances.
   a) oxygen, $O_2$
   b) propane, $C_3H_8$
   c) magnesium sulfate, $MgSO_4$
   d) calcium nitrate, $Ca(NO_3)_2$

2. Sodium reacts with oxygen to make sodium oxide as shown in this equation: $4Na + O_2 \rightarrow 2Na_2O$
   a) Show that the sum of the relative formula masses of all the reactants equals the sum of all the relative formula masses of the products.
   b) In a reaction, 11.5 g of sodium reacts with 4.0 g of oxygen. What mass of sodium oxide is formed?

3. When calcium carbonate is heated it decomposes. In a reaction 10 g of calcium carbonate was heated. At the end of the reaction, only 5.6 g of solid was left. What has happened to the other 4.4 g?

4. What is the mass of one mole of the following substances?
   a) potassium, $K$
   b) nitrogen, $N_2$
   c) sucrose, $C_{12}H_{22}O_{11}$

5. What is the mass of each of the following?
   a) 2.8 moles of chlorine, $Cl_2$
   b) 0.05 moles of methanol, $CH_3OH$

6. How many moles are there in each of the following?
   a) 2.4 g of oxygen, $O_2$
   b) 10 kg of iron oxide, $Fe_2O_3$
   c) 0.5 tonnes of ammonium nitrate, $NH_4NO_3$
   d) 25 mg of platinum, $Pt$

7. a) What mass of water molecules is the same number of molecules as the number of molecules in 88 g of carbon dioxide molecules?
    b) What mass of oxygen molecules is the same number of molecules as the number of atoms in 10 g of calcium atoms?

8. Hydrogen can be made by the reaction of methane with steam. What mass of hydrogen is formed from 80 g of methane?
   $CH_4 + H_2O \rightarrow 3H_2 + CO$

9. Iron for welding railway lines together is produced from the reaction of iron oxide with aluminium. How much iron oxide will react with 1.00 kg of aluminium? Give your answer to 3 significant figures.
   $Fe_2O_3 + 2Al \rightarrow 2Fe + Al_2O_3$

10. When 1.70 g of sodium nitrate ($NaNO_3$) is heated, 1.38 g of sodium nitrite ($NaNO_2$) and 0.32 g of oxygen ($O_2$) are formed. By calculating molar ratios, deduce the balanced equation for this reaction.
11 The salt potassium sulfate can be made by reaction of sulfuric acid with potassium hydroxide.
\[ \text{H}_2\text{SO}_4 + 2\text{KOH} \rightarrow \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O} \]
How many moles of potassium sulfate would be made if:

a) 10 moles of sulfuric acid was reacted with 10 moles of potassium hydroxide?

b) 10 moles of sulfuric acid was reacted with 15 moles of potassium hydroxide?

c) 10 moles of sulfuric acid was reacted with 25 moles of potassium hydroxide?

12 Chromium metal can be made when chromium oxide reacts with aluminium as shown in the equation below. What mass of chromium can be made when 30.4 g of chromium oxide reacts with 13.5 g of aluminium? One of the reagents is in excess. Give your answer to 3 significant figures.
\[ \text{Cr}_2\text{O}_3 + 2\text{Al} \rightarrow 2\text{Cr} + \text{Al}_2\text{O}_3 \]
Practice questions

1 Which one of the following would contain the same number of moles as 6 g of magnesium? [1 mark]
   A 3 g of carbon  B 6 g of carbon  C 20 g of calcium  D 40 g of calcium

2 What is the mass of one mole of calcium nitrate, Ca(NO₃)₂? [1 mark]
   A 82 g  B 164 g  C 204 g  D 220 g

3 Most metals are found naturally in rocks called ores. Some examples are shown in the table:

<table>
<thead>
<tr>
<th>Metal ore</th>
<th>Formula of compound in ore</th>
</tr>
</thead>
<tbody>
<tr>
<td>Galena</td>
<td>PbS</td>
</tr>
<tr>
<td>Haematite</td>
<td>Fe₂O₃</td>
</tr>
<tr>
<td>Dolomite</td>
<td>CaMg(CO₃)₂</td>
</tr>
</tbody>
</table>

- a) Calculate the relative formula mass ($M_r$) of each compound in the metal ore. [3 marks]
- b) The relative formula mass of another metal ore was calculated to be 102. The formula of this ore can be represented as $X_2O_3$. Use this information to calculate the relative atomic mass of metal $X$. Find the identity of metal $X$. [2 marks]
- c) Iron is extracted from the Fe₂O₃ in haematite by reaction with carbon monoxide as shown in the equation below.
  \[ Fe₂O₃ + 3CO \rightarrow 2Fe + 3CO₂ \]
  i) What is meant by the law of conservation of mass? [1 mark]
  ii) By working out the formula masses of the reactants and products in this equation show that the equation follows the law of conservation of mass. [2 marks]

4 Anaemia is a condition that occurs when the body has too few red blood cells. Anaemia often occurs in pregnant women. To prevent anaemia, iron(II) sulfate tablets can be taken to provide the iron needed by the body to produce red blood cells. One brand of iron(II) sulfate tablets contains 200 mg of iron(II) sulfate.
   a) Calculate the relative formula mass ($M_r$) of iron(II) sulfate (FeSO₄). [1 mark]
   b) What is the mass of one mole of iron(II) sulfate? [1 mark]
   c) Calculate the number of moles of iron(II) sulfate present in one tablet. [2 marks]

5 Lead is extracted from the ore galena PbS.
   a) The ore is roasted in air to produce lead(II) oxide PbO. Calculate the maximum mass of lead(II) oxide PbO produced from 4780 g of galena PbS. [3 marks]
   \[ 2PbS + 3O₂ \rightarrow 2PbO + 2SO₂ \]
   b) The lead(II) oxide is reduced to lead by heating it with carbon in a blast furnace. The molten lead is tapped off from the bottom of the furnace.
   \[ PbO + C \rightarrow Pb + CO \]
   Using your answer to part (a) calculate the maximum mass of lead that would eventually be produced. [2 marks]

6 Sodium hydrogen carbonate decomposes when it is heated. 3.36 g of sodium hydrogen carbonate were placed in a test tube and heated in a Bunsen flame for some time.

\[ 2NaHCO₃ \rightarrow Na₂CO₃ + H₂O + CO₂ \]
   a) Calculate the number of moles of sodium hydrogen carbonate used. [2 marks]
   b) Calculate the number of moles of sodium carbonate formed. [1 mark]
   c) Calculate the mass of sodium carbonate expected to be formed. [2 marks]

7 Zinc sulfate crystals are prepared in the laboratory by reacting zinc carbonate with sulfuric acid, as shown in the equation below.

\[ ZnCO₃ + H₂SO₄ \rightarrow ZnSO₄ + H₂O + CO₂ \]
   What is the maximum mass of zinc sulfate which could be formed when 2.5 g of zinc carbonate are reacted with sulfuric acid? [3 marks]

8 Willow bark contains salicylic acid and was once used as a painkiller. Salicylic acid is now used to manufacture aspirin. A student reacted 4.00 g of salicylic acid with 6.50 g of ethanoic anhydride.

\[ C₆H₄(OH)COOH + (CH₃CO)₂O \rightarrow HOOCC₆H₄OCOCH₃ + CH₃COOH \]
   a) How many moles of salicylic acid were used? [2 marks]
   b) How many moles of ethanoic anhydride were present? [2 marks]
   c) What is the maximum number of moles of aspirin which could be formed? [2 marks]
   d) Calculate the maximum mass of aspirin which could be formed. [2 marks]
Units
Many of the calculations used in chemistry will require different units. It is important that you can convert between units.

Volume
Volume is measured in cm³ or dm³ (cubic decimetres) or m³.

1000 cm³ = 1 dm³

You need to be able to convert between these volume units. The flow scheme in Figure 11.16 will help you to convert between volume units.

![Figure 11.16 Converting between volume units.](image)

Example
What is 15 cm³ in dm³?

Answer
To convert from cm³ to dm³ you need to divide by 1000

15 cm³ = \( \frac{15}{1000} = 0.015 \text{ dm}^3 \)

Example
What is 0.4 dm³ in cm³?

Answer
To convert from dm³ to cm³ you need to multiply by 1000

0.4 dm³ = 0.4 × 1000 = 400 cm³

Mass
Mass can be measured in milligrams (mg), grams (g), kilograms (kg) and in tonnes.

1 tonne = 1000 kg
1 kilogram = 1000 g
1 gram = 1000 mg

Questions
1 Convert the following volumes to the units shown.
   a) 25 cm³ to dm³
   b) 100 cm³ to dm³
   c) 10 dm³ to cm³
   d) 20 dm³ to m³
   e) 24 000 cm³ to dm³
The flow diagram in Figure 11.17 will help you to convert between mass units.

\[
\begin{array}{cccc}
\text{tonne} & \times 1000 & \text{kilogram} & \times 1000 \\
\downarrow & & \downarrow & \\
\text{gram} & \div 1000 & \text{milligram} & \div 1000 \\
\end{array}
\]

▲ Figure 11.17 Converting between mass units.

**Example**
Convert 420 mg to grams.

**Answer**
To convert from mg to g you need to divide by 1000:
\[
\frac{420}{1000} = 0.420 \text{ g}
\]

**Example**
Convert 3.2 kg to grams.

**Answer**
To convert from kg to g you need to multiply by 1000:
\[
3.2 \times 1000 = 3200 \text{ g}
\]

**Example**
Convert 0.44 tonnes to grams.

**Answer**
First you need to convert from tonnes to kilograms by multiplying by 1000:
\[
0.44 \times 1000 = 440 \text{ kg}
\]
Then convert 440 kg to g by multiplying by 1000:
\[
440 \times 1000 = 440000 \text{ g}
\]

**Example**
Convert 250 mg to kg.

**Answer**
First you need to convert mg to g by dividing by 1000:
\[
\frac{250}{1000} = 0.25 \text{ g}
\]
Then convert 0.25 g to kg by dividing by 1000:
\[
\frac{0.25}{1000} = 0.00025 = 2.5 \times 10^{-4} \text{ kg}
\]

**Questions**

2 Convert the following masses to the units shown.
   a) 25 g to kg
   b) 1032 kg to tonnes
   c) 10 tonnes to kg
   d) 43 mg to g
   e) 6.13 tonnes to g
   f) 0.3 kg to g

3 Carry out the following unit conversions.
   a) 50 cm\(^3\) to dm\(^3\)
   b) 32000 g to tonnes
   c) 22000 cm\(^3\) to dm\(^3\)
   d) 0.7 kg to g
   e) 2.45 tonnes to g
   f) 12 cm\(^3\) to dm\(^3\)
People have been using chemical reactions to produce metals since the Bronze Age and Iron Age. Chemistry is all about making useful substances from everyday resources and to do this we need to carry out chemical reactions. This chapter looks at some of the most common and important chemical reactions including reactions of metals, reactions of acids and electrolysis.

This chapter covers specification points 5.4.1.1 to 5.4.3.5 and is called Chemical changes. It covers reaction and extraction of metals, reactions of acids, making salts and electrolysis.

Related work on writing equations and half equations can be found in Chapter 14.
Previously you could have learnt:

› Metals can be listed in a reactivity series which compares metals in terms of their reactivity.
› Low reactivity metals include copper, silver and gold.
› High reactivity metals include sodium, calcium and magnesium.
› Metals are extracted from compounds in ores.
› Metals with low reactivity are extracted by heating metal compounds with carbon; high reactivity metals are extracted by electrolysis.
› Metals react with acids to produce hydrogen gas.
› Acids react with carbonates to form carbon dioxide gas.
› Acids have a pH less than 7; neutral solutions have a pH of 7; alkalis have a pH greater than 7.

Test yourself on prior knowledge

1 Name two reactive metals.
2 Name two unreactive metals.
3 Sodium is found in the ore halite which contains sodium chloride. How is sodium extracted from this ore?
4 What gas is made when acids react with:
   a) metals
   b) carbonates?
5 State whether each of the following solutions is acidic, neutral or alkaline.
   a) pH 11
   b) pH 2
   c) pH 7
   d) pH 6

Reactions of metals

The reactivity series of metals

Metals have many uses. For example, they are used in electrical cables, cars, aeroplanes, buildings, mobile phones and computers. Some metals, such as gold, are very unreactive. Other metals, such as sodium, are very reactive.

The reactivity series of metals shows the metals in order of reactivity. This order can be worked out by comparing how metals react with substances such as oxygen, water and dilute acids. The more vigorous the reaction, the higher the reactivity of the metal.

The reactivity series in Figure 12.1 shows some common metals. Carbon and hydrogen are included for comparison although they are non-metals.
Reactions of metals

<table>
<thead>
<tr>
<th>Metal</th>
<th>Reaction with oxygen</th>
<th>Reaction with water</th>
<th>Reaction with acids</th>
</tr>
</thead>
<tbody>
<tr>
<td>K</td>
<td>Burns to form oxide</td>
<td>Reacts and gives off H₂(g)</td>
<td>Reacts violently and gives off H₂(g)</td>
</tr>
<tr>
<td>Na</td>
<td></td>
<td></td>
<td>Reacts and gives off H₂(g)</td>
</tr>
<tr>
<td>Li</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ca</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mg</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Al</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>C</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Zn</td>
<td>Forms oxide when heated</td>
<td>No reaction</td>
<td></td>
</tr>
<tr>
<td>Fe</td>
<td></td>
<td></td>
<td>Reacts slowly and gives off H₂(g)</td>
</tr>
<tr>
<td>Sn</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Pb</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>H</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Cu</td>
<td>Forms oxide when heated</td>
<td>No reaction</td>
<td></td>
</tr>
<tr>
<td>Ag</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Au</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Pt</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Most metals react with oxygen. Reactive metals burn when they are heated as they react with oxygen in the air. The metals in the middle of the reactivity series react with oxygen when heated and can burn if the metal is powdered. Copper, a low reactivity metal, reacts with oxygen forming a layer of copper oxide on the surface of the copper but does not burn (Figure 12.2). Metals with very low reactivity, such as gold, do not react with oxygen at all.

When metals react with oxygen they form a metal oxide:

\[
\text{metal + oxygen \rightarrow metal oxide}
\]

For example:

\[
\text{sodium + oxygen \rightarrow sodium oxide} \quad 4\text{Na} + \text{O}_2 \rightarrow 2\text{Na}_2\text{O}
\]

\[
\text{copper + oxygen \rightarrow copper oxide} \quad 2\text{Cu} + \text{O}_2 \rightarrow 2\text{CuO}
\]

These are examples of oxidation reactions. An oxidation reaction can be defined as a reaction where a substance gains oxygen. A reduction reaction can be defined as a reaction where a substance loses oxygen.
Reaction with water

Most metals do not react with cold water. However, metals with high reactivity react with cold water to form a metal hydroxide and hydrogen gas (Figure 12.3).

\[
\text{metal + water} \rightarrow \text{metal hydroxide} + \text{hydrogen}
\]

Table 12.1 shows how metals react with cold water.

Figure 12.3 Potassium reacts vigorously with water.

<table>
<thead>
<tr>
<th>Metal</th>
<th>Observations</th>
<th>Equation for reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Potassium (K)</td>
<td>Fizzes, melts, floats and moves on the surface of the water, lilac flame (Figure 12.3)</td>
<td>potassium + water $\rightarrow$ potassium hydroxide + hydrogen $2K + 2H_2O \rightarrow 2KOH + H_2$</td>
</tr>
<tr>
<td>Sodium (Na)</td>
<td>Fizzes, melts, floats and moves on the surface of the water (sometimes there is a yellow-orange flame)</td>
<td>sodium + water $\rightarrow$ sodium hydroxide + hydrogen $2Na + 2H_2O \rightarrow 2NaOH + H_2$</td>
</tr>
<tr>
<td>Lithium (Li)</td>
<td>Fizzes, floats and moves on the surface of the water</td>
<td>lithium + water $\rightarrow$ lithium hydroxide + hydrogen $2Li + 2H_2O \rightarrow 2LiOH + H_2$</td>
</tr>
<tr>
<td>Calcium (Ca)</td>
<td>Fizzes, white solid forms</td>
<td>calcium + water $\rightarrow$ calcium hydroxide + hydrogen $Ca + 2H_2O \rightarrow Ca(OH)₂ + H_2$</td>
</tr>
<tr>
<td>Magnesium (Mg)</td>
<td>Very slow reaction</td>
<td>magnesium + water $\rightarrow$ magnesium hydroxide + hydrogen $Mg + 2H_2O \rightarrow Mg(OH)₂ + H_2$</td>
</tr>
<tr>
<td>Zinc (Zn)</td>
<td>No reaction</td>
<td></td>
</tr>
<tr>
<td>Iron (Fe)</td>
<td>No reaction</td>
<td></td>
</tr>
<tr>
<td>Copper (Cu)</td>
<td>No reaction</td>
<td></td>
</tr>
</tbody>
</table>

Reaction with dilute acids

Metals that are more reactive than hydrogen react with dilute acids. When metals react with dilute acids they form a salt and hydrogen gas.

\[
\text{metal + acid} \rightarrow \text{metal salt} + \text{hydrogen}
\]

Hydrochloric acid makes chloride salts. Sulfuric acid makes sulfate salts. Nitric acid makes nitrate salts.

With high reactivity metals, the reaction with acids is explosive due to the hydrogen that is formed igniting. Metals that are less reactive than hydrogen do not react with dilute acids.

Table 12.2 shows how metals react with dilute hydrochloric acid.

Figure 12.4 Magnesium fizzing as it reacts with an acid giving off hydrogen gas.

<table>
<thead>
<tr>
<th>Metal</th>
<th>Observation</th>
<th>Equation for reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Potassium (K)</td>
<td>Explosive</td>
<td>potassium + hydrochloric acid $\rightarrow$ potassium chloride + hydrogen $2K + 2HCl \rightarrow 2KCl + H_2$</td>
</tr>
<tr>
<td>Sodium (Na)</td>
<td>Explosive</td>
<td>sodium + hydrochloric acid $\rightarrow$ sodium chloride + hydrogen $2Na + 2HCl \rightarrow 2NaCl + H_2$</td>
</tr>
<tr>
<td>Lithium (Li)</td>
<td>Explosive</td>
<td>lithium + hydrochloric acid $\rightarrow$ lithium chloride + hydrogen $2Li + 2HCl \rightarrow 2LiCl + H_2$</td>
</tr>
<tr>
<td>Calcium (Ca)</td>
<td>Fizzes</td>
<td>calcium + hydrochloric acid $\rightarrow$ calcium chloride + hydrogen $Ca + 2HCl \rightarrow CaCl₂ + H_2$</td>
</tr>
<tr>
<td>Magnesium (Mg)</td>
<td>Fizzes (Figure 12.4)</td>
<td>magnesium + hydrochloric acid $\rightarrow$ magnesium chloride + hydrogen $Mg + 2HCl \rightarrow MgCl₂ + H₂$</td>
</tr>
<tr>
<td>Zinc (Zn)</td>
<td>Fizzes slowly</td>
<td>zinc + hydrochloric acid $\rightarrow$ zinc chloride + hydrogen $Zn + 2HCl \rightarrow ZnCl₂ + H_2$</td>
</tr>
<tr>
<td>Iron (Fe)</td>
<td>Fizzes slowly</td>
<td>iron + hydrochloric acid $\rightarrow$ iron chloride + hydrogen $Fe + 2HCl \rightarrow FeCl₂ + H_2$</td>
</tr>
<tr>
<td>Copper (Cu)</td>
<td>No reaction</td>
<td></td>
</tr>
</tbody>
</table>
What happens to metal atoms when they react?

When metal atoms react, they lose electrons to form positive ions. For example:

- When sodium atoms react with oxygen, the sodium atoms lose electrons and form sodium ions (Na⁺) in the product sodium oxide.
- When calcium atoms react with water, the calcium atoms lose electrons and form calcium ions (Ca²⁺) in the product calcium hydroxide.
- When zinc atoms react with hydrochloric acid, the zinc atoms lose electrons and form zinc ions (Zn²⁺) in the product zinc chloride.

The greater the tendency of a metal to lose electrons to form ions, the more reactive it is. Reactive metals like potassium and sodium easily lose electrons to form ions, but metals like gold and platinum do not tend to form ions and so are unreactive.

Displacement reactions

In a displacement reaction, a more reactive metal will take the place of a less reactive metal in a compound. For example, aluminium will displace iron from iron oxide because aluminium is more reactive than iron.

\[
\text{aluminium + iron oxide} \rightarrow \text{aluminium oxide} + \text{iron} \\
2\text{Al} + \text{Fe}_2\text{O}_3 \rightarrow \text{Al}_2\text{O}_3 + 2\text{Fe}
\]

This reaction is used to weld railway lines together (Figure 12.5). A mixture of aluminium and iron oxide is placed over the gap between the railway lines and the reaction started. The reaction gets very hot and produces molten iron which flows into a mould, cools and solidifies to weld the lines together.

Displacement reactions also take place in solution. For example, copper will displace silver from silver nitrate solution because copper is more reactive than silver. Copper nitrate is blue and so the solution turns blue as the copper nitrate is formed (Figure 12.6).

\[
\text{copper + silver nitrate} \rightarrow \text{copper nitrate} + \text{silver} \\
\text{Cu} + 2\text{AgNO}_3 \rightarrow \text{Cu(NO}_3)_2 + 2\text{Ag}
\]
Oxidation and reduction in terms of electrons

**Oxidation** can be defined as a reaction where a substance gains oxygen. However, a better definition of oxidation is a reaction where a substance loses electrons.

**Reduction** can be defined as a reaction where a substance loses oxygen. However, a better definition of reduction is a reaction where a substance gains electrons.

---

### Test yourself

1. Complete the following word equations, or write *no reaction* if no reaction would take place.
   a) calcium + oxygen  
   b) gold + oxygen  
   c) copper + water  
   d) lithium + water  
   e) calcium + nitric acid  
   f) copper + sulfuric acid  
   g) zinc + hydrochloric acid  
   h) iron + sulfuric acid  
   i) tin + magnesium chloride  
   j) zinc + lead nitrate  
   k) magnesium + aluminum sulfate

2. Magnesium (Mg) reacts with oxygen (O₂) to form magnesium oxide (MgO).
   a) Write a word equation for this reaction.
   b) Write a balanced equation for this reaction.
   c) What happens to the magnesium atoms in this reaction in terms of electrons?

3. Calcium (Ca) reacts with water to form calcium hydroxide (Ca(OH)₂) and hydrogen (H₂).
   a) Describe what you would see in this reaction.
   b) Write a word equation for this reaction.
   c) Write a balanced equation for this reaction.
   d) What happens to the calcium atoms in this reaction in terms of electrons?

4. Magnesium (Mg) and zinc (Zn) both react with sulfuric acid.
   a) i) Which metal reacts more vigorously with sulfuric acid?  
      ii) Explain, in terms of the tendency to form ions, why this metal reacts more vigorously.
   b) For the reaction of magnesium with sulfuric acid:
      i) Write a word equation  
      ii) Write a balanced equation.

5. The metal chromium can be made in a displacement reaction between aluminium and chromium oxide.
   
   2Al + Cr₂O₃ → Al₂O₃ + 2Cr
   a) Why does aluminium displace chromium in this reaction?
   b) Which substance is oxidised in this reaction?
   c) Which substance is reduced in this reaction?

---

### Show you can...

To determine the order of reactivity of the metals copper, magnesium, nickel and zinc each metal was heated with the oxides of other metals and the results obtained recorded in the table below.

<table>
<thead>
<tr>
<th>Table 12.3</th>
<th>Copper</th>
<th>Magnesium</th>
<th>Nickel</th>
<th>Zinc</th>
</tr>
</thead>
<tbody>
<tr>
<td>Copper oxide</td>
<td>Reaction</td>
<td>Reaction</td>
<td>Reaction</td>
<td></td>
</tr>
<tr>
<td>Magnesium oxide</td>
<td>No reaction</td>
<td>No reaction</td>
<td>No reaction</td>
<td></td>
</tr>
<tr>
<td>Nickel oxide</td>
<td>No reaction</td>
<td>Reaction</td>
<td>Reaction</td>
<td></td>
</tr>
<tr>
<td>Zinc oxide</td>
<td>No reaction</td>
<td>Reaction</td>
<td>No reaction</td>
<td></td>
</tr>
</tbody>
</table>

a) Determine the order of the four metals from the most reactive to the least reactive.
b) Write a balanced chemical equation for the reaction of nickel oxide (NiO) with magnesium.
c) From the list below, write word and balanced chemical equations for all reactions which occur (when nickel reacts, it forms Ni²⁺ ions).
   i) nickel + hydrochloric acid  
   ii) zinc + water  
   iii) nickel + water  
   iv) zinc + sulfuric acid  
   v) magnesium + zinc oxide

---

### KEY TERMS

**Oxidation** A reaction in which a substance loses electrons (gains oxygen).

**Reduction** Reaction in which a substance gains electrons (loses oxygen).
Reactions of metals

Table 12.4 Reactions of metals.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>General equation</th>
<th>Oxidation in terms of gaining oxygen</th>
<th>Oxidation in terms of losing electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>+ oxygen</td>
<td>metal + oxygen → metal oxide</td>
<td>Metal gains oxygen</td>
<td>Metal atoms lose electrons to form metal ions in the metal oxide</td>
</tr>
<tr>
<td>+ water</td>
<td>metal + water → metal hydroxide + hydrogen</td>
<td>Metal atoms lose electrons to form metal ions in the metal hydroxide</td>
<td>Metal atoms lose electrons to form metal ions in the metal hydroxide</td>
</tr>
<tr>
<td>+ acids</td>
<td>metal + acid → metal salt + hydrogen</td>
<td>Metal atoms lose electrons to form metal ions in the metal salt</td>
<td>Metal atoms lose electrons to form metal ions in the metal salt</td>
</tr>
</tbody>
</table>

Displacement reactions involve oxidation and reduction. In some reactions this can be explained in terms of oxygen and in terms of electrons. This is the case, for example, in the displacement of iron from iron oxide by aluminium (Figure 12.8).

In other displacement reactions, only the definitions in terms of electrons can be used to define the reaction as involving oxidation and reduction. This is the case, for example, in the displacement of silver from silver nitrate by copper (Figure 12.9).

In a reaction in which one substance loses electrons, another substance gains those electrons. This means that both reduction and oxidation take place and these are called redox reactions (reduction–oxidation).
Writing ionic equations and/or half equations for displacement reactions

Ionic equations and/or half equations can be written for displacement reactions. (See Chapter 14 for more help on writing these equations.)

**Example**

Write two half equations for the displacement of iron from iron oxide by aluminium.

\[
\text{aluminium + iron oxide} \to \text{aluminium oxide + iron}
\]

\[
2\text{Al} + \text{Fe}_2\text{O}_3 \to \text{Al}_2\text{O}_3 + 2\text{Fe}
\]

**Answer**

In this reaction, the Al atoms become Al\(^{3+}\) ions in Al\(_2\)O\(_3\) while the Fe\(^{3+}\) ions in Fe\(_2\)O\(_3\) become Fe atoms.

The two half equations for this are:

- Al atoms lose electrons to form Al\(^{3+}\) ions: Al \(-\text{3e}^-\to\) Al\(^{3+}\) (or Al \(-\text{Al}^{3+} + 3\text{e}^-\))
- Fe\(^{3+}\) ions in Fe\(_2\)O\(_3\) gain electrons to form Fe atoms: Fe\(^{3+}\) + 3e\(^-\) \to Fe

**Example**

Write an overall ionic equation and two half equations for the displacement of silver from silver nitrate by copper.

\[
copper + silver nitrate \to copper nitrate + silver
\]

\[
\text{Cu} + 2\text{AgNO}_3 \to \text{Cu(NO}_3\text{)}_2 + 2\text{Ag}
\]

**Answer**

In this reaction, the Cu atoms become Cu\(^{2+}\) ions in Cu(NO\(_3\))\(_2\) while the Ag\(^+\) ions in AgNO\(_3\) become Ag atoms. We can leave out the NO\(_3^-\) ions from the ionic equation as they do not change.

Therefore, the overall ionic equation is: Cu + 2Ag\(^+\) \to Cu\(^{2+}\) + 2Ag

The two half equations for this are: Cu atoms lose electrons to form Cu\(^{2+}\) ions: Cu \(-\text{2e}^-\to\) Cu\(^{2+}\) (or Cu \(-\text{Cu}^{2+} + 2\text{e}^-\))

- Ag\(^+\) ions in AgNO\(_3\) gain electrons to form Ag atoms: Ag\(^+\) + e\(^-\) \to Ag

**Example**

Write an overall ionic equation and two half equations for the displacement of copper from copper sulfate by iron. Identify which half equation represents a reduction process and which represents an oxidation process.

\[
iron + copper sulfate \to iron sulfate + copper
\]

\[
\text{Fe} + \text{CuSO}_4 \to \text{FeSO}_4 + \text{Cu}
\]

**Answer**

In this reaction, the Fe atoms become Fe\(^{2+}\) ions in FeSO\(_4\) while the Cu\(^{2+}\) ions in CuSO\(_4\) become Cu atoms. We can leave out the SO\(_4^-\) ions from the ionic equation as they do not change.

Therefore, the overall ionic equation is: Fe + Cu\(^{2+}\) \to Fe\(^{2+}\) + Cu

The two half equations for this are:

- Fe atoms lose electrons to form Fe\(^{2+}\) ions: Fe \(-\text{2e}^-\to\) Fe\(^{2+}\) (or Fe \(-\text{Fe}^{2+} + 2\text{e}^-\))
- Cu\(^{2+}\) ions in CuSO\(_4\) gain electrons to form Cu atoms: Cu\(^{2+}\) + 2e\(^-\) \to Cu

The reduction half equation is: Cu\(^{2+}\) + 2e\(^-\) \to Cu

The oxidation half equation is: Fe \(-\text{2e}^-\to\) Fe\(^{2+}\) (or Fe \(-\text{Fe}^{2+} + 2\text{e}^-\))
Test yourself

6 Rubidium is a metal in Group 1 of the periodic table. It reacts vigorously with water.
   a) Write a word equation for this reaction.
   b) Write a balanced equation for this reaction.
   c) What happens to the rubidium atoms in this reaction in terms of electrons?
   d) Are the rubidium atoms oxidised or reduced in this reaction? Explain your answer.

7 Magnesium reacts with copper oxide in a displacement reaction to form copper.
   \( \text{Mg} + \text{CuO} \rightarrow \text{MgO} + \text{Cu} \)
   a) Explain, in terms of oxygen, why the magnesium is oxidised in this reaction.
   b) Write a half equation to show the oxidation of magnesium atoms to magnesium ions in this reaction.
   c) Explain, in terms of electrons, why the magnesium is oxidised in this reaction.
   d) Explain, in terms of oxygen, why the copper oxide is reduced in this reaction.
   e) Write a half equation to show the reduction of copper ions to copper atoms in this reaction.
   f) Explain, in terms of electrons, why the copper oxide is reduced in this reaction.

8 Write an overall ionic equation and two half equations for each of the following displacement reactions.
   a) Displacement of copper from copper sulfate by zinc.
      \( \text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu} \)
   b) Displacement of silver from silver nitrate by zinc.
      \( \text{Zn} + 2\text{AgNO}_3 \rightarrow \text{Zn(NO}_3)_2 + 2\text{Ag} \)
   c) Displacement of copper from copper sulfate by aluminium.
      \( 2\text{Al} + 3\text{CuSO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + 3\text{Cu} \)

9 Zinc displaces iron from a solution of iron(II) sulfate. For this reaction:
   a) Write a word equation for this reaction.
   b) Write a balanced equation for this reaction.
   c) Write an ionic equation for this reaction.
   d) Write the two half equations for this reaction.
   e) Identify which half equation is a reduction process.
   f) Identify which half equation is an oxidation process.
   g) Explain why this is a redox reaction.

10 Magnesium displaces silver from a solution of silver nitrate. For this reaction:
   a) Write a word equation for this reaction.
   b) Write a balanced equation for this reaction.
   c) Write an ionic equation for this reaction.
   d) Write the two half equations for this reaction.
   e) Identify which half equation is a reduction process.
   f) Identify which half equation is an oxidation process.
   g) Explain why this is a redox reaction.
12 Chemical changes

Extraction of metals

Where do metals come from?

A few metals, such as gold and platinum, occur naturally on Earth as elements. These are metals with very low reactivity.

Most metals are only found on Earth in compounds. For example, iron is often found in the compound iron oxide and aluminium in the compound aluminium oxide. In order to extract the metal from these compounds, a chemical reaction is required.

The metal compounds are found in rocks called ores. An ore is a rock from which a metal can be extracted for profit (Figure 12.10).

Methods of extraction

Most of the compounds from which metals are extracted are oxides. In order to extract the metal, the oxygen is removed in a reduction reaction. The way in which this is done depends on the reactivity of the metal (Figure 12.11).

- Metals that are less reactive than carbon can be extracted by heating the metal oxide with carbon. For example, iron is extracted by heating iron oxide with carbon. The iron oxide is reduced in this reaction (Figure 12.12).
- Metals that are more reactive than carbon can be extracted by electrolysis. This is studied in detail in the section on electrolysis later in this chapter.

When metals are extracted from compounds, the metal ions in the compounds gain electrons. For example, when iron is extracted from iron oxide, Fe$^{3+}$ ions in the iron oxide gain electrons to form Fe atoms. This is reduction because the iron oxide loses oxygen and also because the Fe$^{3+}$ ions in the iron oxide gain electrons.

<table>
<thead>
<tr>
<th>Metal</th>
<th>Method of extraction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Potassium</td>
<td>Electrolysis</td>
</tr>
<tr>
<td>Sodium</td>
<td>Electrolysis</td>
</tr>
<tr>
<td>Lithium</td>
<td>Electrolysis</td>
</tr>
<tr>
<td>Calcium</td>
<td>Electrolysis</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Electrolysis</td>
</tr>
<tr>
<td>Aluminium</td>
<td>Electrolysis</td>
</tr>
<tr>
<td>Carbon</td>
<td>Heat with carbon</td>
</tr>
<tr>
<td>Zinc</td>
<td>Heat with carbon</td>
</tr>
<tr>
<td>Iron</td>
<td>Heat with carbon</td>
</tr>
<tr>
<td>Tin</td>
<td>Heat with carbon</td>
</tr>
<tr>
<td>Lead</td>
<td>Heat with carbon</td>
</tr>
<tr>
<td>Copper</td>
<td>Heat with carbon</td>
</tr>
<tr>
<td>Silver</td>
<td>Heat with carbon</td>
</tr>
<tr>
<td>Gold</td>
<td>Heat with carbon</td>
</tr>
<tr>
<td>Platinum</td>
<td>Heat with carbon</td>
</tr>
</tbody>
</table>

KEY TERM

Ore A rock from which a metal can be extracted for profit.

\[
\text{Fe}_2\text{O}_3 + 3\text{C} \rightarrow 2\text{Fe} + 3\text{CO}
\]

- Figure 12.10 Aluminium metal is extracted from the aluminium oxide (Al$_2$O$_3$) in the ore bauxite.

- Figure 12.11 Extracting metals from their oxides.

- Figure 12.12 Extracting iron from iron oxide by heating with carbon.
Reactions of acids

**What are acids and alkalis?**

An **acid** is a substance that produces hydrogen ions, \( \text{H}^+ \), in **aqueous** solution. For example, solutions of:

- hydrochloric acid (HCl) contain hydrogen (\( \text{H}^+ \)) ions and chloride (\( \text{Cl}^- \)) ions
- sulfuric acid (\( \text{H}_2\text{SO}_4 \)) contain hydrogen (\( \text{H}^+ \)) ions and sulfate (\( \text{SO}_4^{2-} \)) ions
- nitric acid (\( \text{HNO}_3 \)) contain hydrogen (\( \text{H}^+ \)) ions and nitrate (\( \text{NO}_3^- \)) ions.

An **alkali** is a substance that produces hydroxide (\( \text{OH}^- \)) ions, in aqueous solution. For example, solutions of:

- sodium hydroxide (\( \text{NaOH} \)) contain sodium (\( \text{Na}^+ \)) ions and hydroxide (\( \text{OH}^- \)) ions
- potassium hydroxide (\( \text{KOH} \)) contain potassium (\( \text{K}^+ \)) ions and hydroxide (\( \text{OH}^- \)) ions
- calcium hydroxide (\( \text{Ca(OH)}_2 \)) contain calcium (\( \text{Ca}^{2+} \)) ions and hydroxide (\( \text{OH}^- \)) ions.

**The pH scale**

The pH scale is a measure of how acidic or alkaline a solution is. A solution with a pH of 7 is neutral, whereas a solution with a pH below 7 is acidic and one with a pH above 7 is alkaline. The further away from 7 the pH is, the more acidic or alkaline the solution is (Figure 12.13).

The scale is often shown as running from 0 to 14, but it does go further in both directions. For example, it is common for the solutions of acids...
12 Chemical changes

in school laboratories to have a pH that is less than 0 (typically about −0.3).

The approximate pH of a solution can be measured using universal indicator solution. A few drops of the indicator is added to the solution. The colour is compared to a colour chart to give the approximate pH of the solution (Figure 12.14).

A more accurate way of finding the pH of a solution is to use a pH probe (Figure 12.15). There are different types but the probe is dipped into the solution and the pH shown on the display, often to 1 or 2 decimal places.

The pH of a solution is based on the concentration of H\(^+\) ions in the solution. The higher the concentration of H\(^+\) ions the lower the pH. As the pH decreases by one unit, the concentration of hydrogen ions increases by a factor of 10. For example, a solution with a pH of 2 has a concentration of H\(^+\) ions that is 10 times greater than one with a pH of 3. A solution with a pH of 1 has a concentration of H\(^+\) ions that is 100 times greater than one with a pH of 3.

In a neutral solution, the concentration of H\(^+\) ions equals the concentration of OH\(^-\) ions.

**Show you can...**

In an experiment a sample of human saliva was removed from the mouth every five minutes after a meal and the pH value determined. The graph shows how the pH value of the saliva changed.

**Test yourself**

14 Classify each of the following solutions as acidic, neutral or alkaline.
   a) A solution with pH 9.
   b) A solution with pH 3.
   c) A solution with pH 0.
   d) A solution with pH 7.

15 a) Three solutions had pH values of 8, 11 and 13. Which one was the most alkaline?
   b) Three solutions had pH values of 1, 2 and 5. Which one was the most acidic?

16 a) Give two ways in which the pH of a solution can be measured.
   b) Which method will give the most accurate value?

17 a) Which ion do aqueous solutions of acids all contain?
   b) Which ion do aqueous solutions of alkalis all contain?

18 The table gives some information about three solutions.

<table>
<thead>
<tr>
<th>Table 12.5</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solution</td>
</tr>
<tr>
<td>pH</td>
</tr>
</tbody>
</table>

a) Which solution has the highest concentration of H\(^+\) ions?
   b) By what factor is concentration of H\(^+\) ions in solution B bigger or smaller than solution A?
   c) By what factor is concentration of H\(^+\) ions in solution C bigger or smaller than solution A?

19 A solution has a pH of 7. Comment on the concentration of H\(^+\) ions compared to the concentration of OH\(^-\) ions.
### Strong and weak acids

Hydrogen chloride (HCl), hydrogen sulfate (H₂SO₄) and hydrogen nitrate (HNO₃) are molecules in their pure state. When they are added to water all of their molecules break down into ions forming hydrochloric acid (HCl), sulfuric acid (H₂SO₄) and nitric acid (HNO₃). They are strong acids because their molecules are completely ionised in water – this means that all their molecules break into ions in water (Figure 12.17).

\[
\begin{align*}
\text{(a)} & \quad \text{HCl(g)} + \text{H₂O} \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \\
\text{(b)} & \quad \text{H₂SO₄(l)} + \text{H₂O} \rightarrow 2\text{H}^+(\text{aq}) + \text{SO₄}^{2-}(\text{aq})
\end{align*}
\]

In weak acids the molecules are only partially ionised in water – this means that only a small fraction of the molecules break into ions when added to water. Figure 12.18 shows the difference between the strong acid HCl and a weak acid HX in water.

**Figure 12.18** The difference between strong acid (HCl) and a weak acid (HX) in water.

The results of equal concentration of a strong acid and weak acid are compared, there will be more H⁺ ions in the strong acid solution. This means that the solution of the strong acid will have a lower pH.

There are many weak acids in food and drink. They tend to have a sour taste and are not dangerous because there is only a low concentration of H⁺ ions. Ethanoic acid in vinegar, citric acid in citrus fruits (Figure 12.19) and carbonic acid in fizzy drinks are examples of weak acids in food and drink.

### KEY TERMS

**Strong acid** Acid in which all the molecules break into ions in water.

**Weak acid** Acid in which only a small fraction of the molecules break into ions in water.

**Figure 12.17** (a) Hydrogen chloride reacts with water and breaks up into ions. (b) Hydrogen sulfate reacts with water and breaks up into ions.

**Figure 12.18** The difference between strong acid (HCl) and a weak acid (HX) in water.

**Figure 12.19** Citrus fruits contain citric acid.
The terms strong and weak refer to the degree of ionisation in water of acids. The terms dilute and concentrated refer to the amount of acid dissolved in the solution. This is summarised in Figure 12.20.

**Figure 12.20** Differences between strong and weak acids, and concentrated and dilute acids.

---

**KEY TERMS**

**Dilute** A solution in which there is a small amount of solute dissolved.

**Concentrated** A solution in which there is a lot of solute dissolved.

---

**Test yourself**

20 a) Sulfuric acid is a strong acid.

b) Citric acid is a weak acid.

---

**Show you can...**

Two solutions A and B were tested using a pH meter, red litmus paper, blue litmus paper, and universal indicator paper. The results are shown in the table.

<table>
<thead>
<tr>
<th>Test</th>
<th>Result for solution A</th>
<th>Result for solution B</th>
</tr>
</thead>
<tbody>
<tr>
<td>pH meter</td>
<td>1.82</td>
<td>3.85</td>
</tr>
<tr>
<td>Red litmus</td>
<td>Red</td>
<td>Red</td>
</tr>
<tr>
<td>Blue litmus</td>
<td>Red</td>
<td>Red</td>
</tr>
<tr>
<td>Universal indicator paper</td>
<td>Red</td>
<td>Orange</td>
</tr>
</tbody>
</table>

a) Describe how the results with universal indicator may be converted into a pH value.

b) Are solutions A and B acidic, neutral or alkaline?

c) If the experiment was repeated using a more concentrated solution of A would the results be different? Explain your answer.

d) Explain why universal indicator gives more information than litmus.
Reactions of acids

Metals that are more reactive than hydrogen react with dilute acids. When metals react with dilute acids, they form a salt and hydrogen gas. The reaction fizzes as hydrogen gas is produced.

\[
\text{metal + acid} \rightarrow \text{metal salt + hydrogen}
\]

Table 12.7 shows how different acids form different types of salts. Table 12.8 shows how magnesium, zinc and iron react with dilute hydrochloric and sulfuric acids.

A redox reaction takes place when a metal reacts with an acid. Figure 12.21 shows a redox reaction where magnesium reacts with hydrochloric acid.

Table 12.7 Salts formed when magnesium reacts with different acids.

<table>
<thead>
<tr>
<th>Acid</th>
<th>Reaction Type</th>
<th>Type of salt formed</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrochloric acid</td>
<td>magnesium + hydrochloric acid → magnesium chloride + hydrogen</td>
<td>Chloride (contains Cl(^-) ions)</td>
</tr>
<tr>
<td>Sulfuric acid</td>
<td>magnesium + sulfuric acid → magnesium sulfate + hydrogen</td>
<td>Sulfate (contains SO(_4^{2-}) ions)</td>
</tr>
<tr>
<td>Nitric acid</td>
<td>magnesium + nitric acid → magnesium nitrate + hydrogen</td>
<td>Nitrate (contains NO(_3^{-}) ions)</td>
</tr>
</tbody>
</table>

Table 12.8 Reactions between some metals and acids.

<table>
<thead>
<tr>
<th>Metal</th>
<th>Reaction with hydrochloric acid</th>
<th>Reaction with sulfuric acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>Magnesium</td>
<td>Fizzes vigorously</td>
<td>Fizzes vigorously</td>
</tr>
<tr>
<td>Zinc</td>
<td>Fizzes gently</td>
<td>Fizzes gently</td>
</tr>
<tr>
<td>Iron</td>
<td>Fizzes very slowly</td>
<td>Fizzes very slowly</td>
</tr>
</tbody>
</table>

Reactions of acids with metal hydroxides

Acids react with metal hydroxides to form a salt and water. Some examples of this reaction are shown in the Table 12.9.

\[
\text{metal hydroxide + acid} \rightarrow \text{metal salt + water}
\]

The ionic equation for each of these reactions is the same. In each reaction, H\(^+\) ions from the acid are reacting with OH\(^-\) ions from the metal hydroxide. This produces water in each case.

Table 12.9 Reactions between some metal hydroxides and acids.

<table>
<thead>
<tr>
<th>Metal</th>
<th>Reaction with nitric acid</th>
<th>Reaction with hydrochloric acid</th>
<th>Reaction with sulfuric acid</th>
<th>Reaction with hydrochloric acid</th>
<th>Reaction with sulfuric acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium hydroxide</td>
<td>NaOH + HNO(_3) \rightarrow NaNO(_3) + H(_2)O</td>
<td>NaOH + 2HCl \rightarrow NaCl + H(_2)</td>
<td>NaOH + H(_2)SO(_4) \rightarrow NaSO(_4) + H(_2)O</td>
<td>NaOH + 2HCl \rightarrow NaCl + H(_2)</td>
<td>NaOH + H(_2)SO(_4) \rightarrow NaSO(_4) + H(_2)O</td>
</tr>
<tr>
<td>Potassium hydroxide</td>
<td>K(_2)OH + HNO(_3) \rightarrow K(_2)NO(_3) + H(_2)O</td>
<td>K(_2)OH + 2HCl \rightarrow KCl + H(_2)</td>
<td>K(_2)OH + H(_2)SO(_4) \rightarrow KSO(_4) + 2H(_2)O</td>
<td>K(_2)OH + 2HCl \rightarrow KCl + H(_2)</td>
<td>K(_2)OH + H(_2)SO(_4) \rightarrow KSO(_4) + 2H(_2)O</td>
</tr>
<tr>
<td>Calcium hydroxide</td>
<td>Ca(OH)(_2) + HNO(_3) \rightarrow Ca(NO(_3)) + H(_2)O</td>
<td>Ca(OH)(_2) + 2HCl \rightarrow CaCl(_2) + H(_2)</td>
<td>Ca(OH)(_2) + H(_2)SO(_4) \rightarrow CaSO(_4) + 2H(_2)O</td>
<td>Ca(OH)(_2) + 2HCl \rightarrow CaCl(_2) + H(_2)</td>
<td>Ca(OH)(_2) + H(_2)SO(_4) \rightarrow CaSO(_4) + 2H(_2)O</td>
</tr>
</tbody>
</table>
Metal hydroxides that dissolve in water are alkalis because they release OH⁻ ions into the water. Metal hydroxides that are insoluble in water do not release OH⁻ ions into the water and so are not alkalis. However, they still react with acids to produce a salt and water.

In these reactions, the H⁺ ions from the acid are being used up and so they are examples of neutralisation reactions.

○ **Reaction of acids with metal oxides**

Acids react with metal oxides to form a salt and water. Most metal oxides are insoluble in water and the reactions usually need to be heated. Some examples of this reaction are shown in Table 12.10.

\[
\text{metal oxide} + \text{acid} \rightarrow \text{metal salt} + \text{water}
\]

<table>
<thead>
<tr>
<th>Full equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>CaO + 2HNO₃ → Ca(NO₃)₂ + H₂O</td>
</tr>
<tr>
<td>CuO + H₂SO₄ → CuSO₄ + H₂O</td>
</tr>
<tr>
<td>Li₂O + 2HCl → 2LiCl + H₂O</td>
</tr>
</tbody>
</table>

In these reactions, the H⁺ ions from the acid are being used up and so they are examples of neutralisation reactions.

○ **Reaction of acids with metal carbonates**

Acids react with metal carbonates to form a salt, water and carbon dioxide. These reactions usually take place readily. The reaction fizzes as carbon dioxide gas is produced. Some examples of this reaction are shown in the Table 12.11.

\[
\text{metal carbonate} + \text{acid} \rightarrow \text{metal salt} + \text{water} + \text{carbon dioxide}
\]

<table>
<thead>
<tr>
<th>Full equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na₂CO₃ + 2HNO₃ → 2NaNO₃ + H₂O + CO₂</td>
</tr>
<tr>
<td>CuCO₃ + H₂SO₄ → CuSO₄ + H₂O + CO₂</td>
</tr>
<tr>
<td>CaCO₃ + 2HCl → CaCl₂ + H₂O + CO₂</td>
</tr>
</tbody>
</table>

In these reactions, the H⁺ ions from the acid are being used up and so they are examples of neutralisation reactions.
Test yourself

21 Complete the following word equations.
   a) calcium + hydrochloric acid
   b) tin + sulfuric acid
   c) barium hydroxide + nitric acid
   d) lithium hydroxide + hydrochloric acid
   e) nickel oxide + nitric acid
   f) magnesium oxide + sulfuric acid
   g) potassium carbonate + nitric acid
   h) zinc carbonate + hydrochloric acid

22 a) There is fizzing when hydrochloric acid reacts with calcium. What gas causes this?
   b) There is fizzing when hydrochloric acid reacts with calcium carbonate. What gas causes this?

23 Hydrochloric acid reacts with sodium hydroxide as shown in this equation:
   \[ \text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)} \]
   a) What does the symbol (aq) mean?
   b) Write an ionic equation for this reaction.
   c) Explain why sodium hydroxide is an alkali.

24 Write a balanced equation for each of the reactions in question 22(a), (d), (f) and (g).

Show you can...

Sodium sulfate, produced by the reaction between sulfuric acid and sodium hydroxide, is used in washing powders.

a) Write a word equation and a balanced symbol equation for the reaction of sulfuric acid with sodium hydroxide.

The composition of two washing powders, A and B are shown below.

Table 12.12

<table>
<thead>
<tr>
<th>Percentage composition in %</th>
<th>Sodium sulfate</th>
<th>Sodium carbonate</th>
<th>Sodium silicate</th>
<th>Soap</th>
<th>Detergent</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>29</td>
<td>20</td>
<td>20</td>
<td>0</td>
<td>15</td>
</tr>
<tr>
<td>B</td>
<td>35</td>
<td>0</td>
<td>26</td>
<td>6</td>
<td>13</td>
</tr>
</tbody>
</table>

Dilute nitric acid was added to each of the powders. Only one of the powders reacted.

b) Which powder reacted A or B? Explain your answer.

c) During the reaction, effervescence was noted. Name the gas produced in the reaction.

d) Describe a chemical test for this gas and state the result for a positive test.

e) The silicate ion is \( \text{SiO}_3^{2-} \). Suggest the formula for sodium silicate.
Making salts

What are salts?

Salts are substances made when acids react with metals, metal hydroxides, metal oxides and metal carbonates.

Salts are very useful substances. For example:

- many medicines are salts (Figure 12.22)
- fertilisers are salts
- toothpaste contains salts
- many food additives are salts.

Salts are made up of a metal ion combined with the ion left over from the acid when the $H^+$ ions react. The salt produced depends on which acid is used and what it reacts with. Table 12.13 below gives some examples of the salts formed when acids react.

<table>
<thead>
<tr>
<th>Hydrochloric acid</th>
<th>Sulfuric acid</th>
<th>Nitric acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>$HCl$</td>
<td>$H_2SO_4$</td>
<td>$HNO_3$</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Magnesium chloride ($+H_2$)</td>
<td>Magnesium sulfate ($+H_2$)</td>
</tr>
<tr>
<td>$Mg$</td>
<td>$MgCl_2$</td>
<td>$MgSO_4$</td>
</tr>
<tr>
<td>Sodium hydroxide</td>
<td>Sodium chloride ($+H_2O$)</td>
<td>Sodium sulfate ($+H_2O$)</td>
</tr>
<tr>
<td>$NaOH$</td>
<td>$NaCl$</td>
<td>$Na_2SO_4$</td>
</tr>
<tr>
<td>Copper oxide</td>
<td>Copper chloride ($+H_2O$)</td>
<td>Copper sulfate ($+H_2O$)</td>
</tr>
<tr>
<td>$CuO$</td>
<td>$CuCl_2$</td>
<td>$CuSO_4$</td>
</tr>
<tr>
<td>Zinc carbonate</td>
<td>Zinc chloride ($+H_2O + CO_2$)</td>
<td>Zinc sulfate ($+H_2O + CO_3$)</td>
</tr>
<tr>
<td>$ZnCO_3$</td>
<td>$ZnCl_2$</td>
<td>$ZnSO_4$</td>
</tr>
</tbody>
</table>

Making soluble salts

It is important to be able to make pure samples of salts, especially if they are being used in medicines or food.

Some salts are soluble in water and some are not. Figure 12.23 shows the method for making salts that are soluble in water.

It is easier to make a pure salt by reacting an acid with a substance that is insoluble in water rather than one that is soluble. Suitable substances to use include some metals, metal oxides, metal hydroxides or metal carbonates. As it is insoluble, you can add an excess of that substance to ensure that all the acid is used up. The excess can then be filtered off. In this way, there is no left over acid or the substance it reacts with mixed in with the salt that is formed.

After filtration, you are left with a solution of the salt in water. If some of the water is boiled off, a hot saturated solution is formed. As this hot, saturated solution cools, crystals form as the salt is less soluble at lower temperatures and so cannot all stay dissolved.
Test yourself

25 Suggest two chemicals that could be reacted together to make the following salts.
   a) calcium chloride
   b) copper nitrate
   c) aluminium sulfate

26 The salt iron(II) sulfate is used in iron tablets. It can be made by reacting an excess of iron with sulfuric acid.
   a) Why is it important that the iron(II) sulfate used in iron tablets is pure?
   b) Why is an excess of iron used?
   c) How is the excess iron removed?
   d) Write a word equation for the reaction between iron and sulfuric acid.
   e) Write a balanced equation for the reaction between iron and sulfuric acid.

27 Crystals of the salt nickel nitrate form as a hot, saturated solution of nickel nitrate cools down.
   a) What is a saturated solution?
   b) Explain why crystals of nickel nitrate form as the hot, saturated solution cools down.

28 The salt copper chloride can be made by reacting hydrochloric acid with copper oxide, copper carbonate or copper hydroxide. It cannot be made by reacting copper with hydrochloric acid.
   a) Explain why copper chloride cannot be made by reacting copper with hydrochloric acid.
   b) For each of the reactions of hydrochloric acid with copper oxide, copper carbonate and copper hydroxide:
      i) Write a word equation.
      ii) Write a balanced equation.

Show you can...

A solution of the salt magnesium chloride can be prepared by any of reactions A to D in the diagram.

![Diagram](image)

a) Write word equations for each of the reactions A to D.

b) State two observations that you would make during reaction D.
Preparation of a pure dry sample of a soluble salt from an insoluble oxide or carbonate (magnesium sulfate from magnesium carbonate)

Bath crystals are a mixture of water soluble solids which are added to bathwater for health benefits. Bath crystals contain Epsom salts (Figure 12.25) (hydrated magnesium sulfate) which relax muscles, reduce inflammation and help muscle function. The name Epsom salts comes from the town of Epsom, which has mineral springs from which hydrated magnesium sulfate was extracted.

To produce pure dry crystals of hydrated magnesium sulfate (MgSO₄·7H₂O) in the laboratory by reacting magnesium carbonate with sulfuric acid, the following method was followed.

- Measure 25 cm³ of dilute sulfuric acid and place in a conical flask.
- Warm the dilute sulfuric acid using a Bunsen burner and add magnesium carbonate, stirring until it is in excess.
- Filter the solution.
- Heat the filtered solution to make it more concentrated.
- Cool and crystallise.
- Filter the crystals from the solution.
- Dry the crystals.

Questions
1. What piece of apparatus would you use to measure 25 cm³ of sulfuric acid?
2. Draw a labelled diagram of the apparatus used for the second step of the method.
3. Explain why the magnesium carbonate is added until it is in excess.
4. State one way in which you would know that the magnesium carbonate is in excess.
5. What is the general name given to the solid trapped by the filter paper?
6. Why was the filtered solution evaporated using a water bath or electric heater?
7. Why is the solution not evaporated to dryness?
8. Why do crystals form as the solution is cooled?
10. Write a word and balanced symbol equation for the reaction between magnesium carbonate and sulfuric acid.

Electrolysis

What is electrolysis?

Electrolysis is the decomposition of ionic compounds using electricity. Ionic compounds contain metals combined with non-metals. Examples include

- sodium chloride (NaCl) – a combination of the metal sodium with the non-metal chlorine
- copper sulfate (CuSO₄) – a combination of the metal copper with the non-metals sulfur and oxygen.

Ionic compounds are made up of positive and negative ions. As solids, ionic compounds cannot conduct electricity because the ions cannot...
move around. However, when ionic compounds are melted or dissolved, the ions are free to move and conduct electricity. These liquids or solutions are called electrolytes because they are able to conduct electricity.

If two electrodes connected to a supply of electricity are put into the electrolyte, the negative ions are attracted to the positive electrode (anode) and the positive ions are attracted to the negative electrode (cathode) (Figure 12.26). This happens because opposite charges attract each other.

When the ions reach the electrodes they are discharged. This means that they gain or lose electrons so that they lose their charge and become neutral. Positive ions gain electrons. Negative ions lose electrons.

The negative ions are discharged by losing electrons at the positive electrode. These electrons move around the circuit through the wires to the negative electrode. The positive ions are discharged by gaining electrons at the negative electrode.

Table 12.14 Electrodes and electrolysis.

<table>
<thead>
<tr>
<th>Electrode</th>
<th>+ Electrode (anode)</th>
<th>– Electrode (cathode)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Which ions are attracted to the electrode</td>
<td>– ions</td>
<td>+ ions</td>
</tr>
<tr>
<td>What happens at the electrodes</td>
<td>– ions are discharged by losing electrons</td>
<td>+ ions are discharged by gaining electrons</td>
</tr>
<tr>
<td>Oxidation or reduction</td>
<td>Oxidation (loss of electrons)</td>
<td>Reduction (gain of electrons)</td>
</tr>
</tbody>
</table>

Electrolysis of molten ionic compounds

Binary ionic compounds are ones made from one metal combined with one non-metal. Examples include lead bromide (PbBr₂) (Figure 12.27), sodium chloride (NaCl) and aluminium oxide (Al₂O₃) (Table 12.15). The electrolysis of these compounds when molten produces the metal and non-metal.

The metal is produced at the cathode and the non-metal at the anode.

Table 12.15 Products from the electrolysis of aqueous solutions of ionic compounds.

<table>
<thead>
<tr>
<th>Electrode</th>
<th>Ions</th>
<th>+ Electrode (anode)</th>
<th>– Electrode (cathode)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lead bromide, PbBr₂</td>
<td>Lead ions, Pb²⁺, bromide ions, Br⁻</td>
<td>Bromide ions (Br⁻) lose electrons to form bromine (Br₂)</td>
<td>Lead ions (Pb²⁺) gain electrons to form lead (Pb)</td>
</tr>
<tr>
<td></td>
<td></td>
<td>2Br⁻ − 2e⁻ → Br₂</td>
<td>Pb²⁺ + 2e⁻ → Pb</td>
</tr>
<tr>
<td>Sodium chloride, NaCl</td>
<td>Sodium ions, Na⁺, chloride ions, Cl⁻</td>
<td>Chloride ions (Cl⁻) lose electrons to form chlorine (Cl₂)</td>
<td>Sodium ions (Na⁺) gain electrons to form sodium (Na)</td>
</tr>
<tr>
<td></td>
<td></td>
<td>2Cl⁻ − 2e⁻ → Cl₂</td>
<td>Na⁺ + e⁻ → Na</td>
</tr>
<tr>
<td>Aluminium oxide, Al₂O₃</td>
<td>Aluminium ions, Al³⁺, oxide ions, O₂⁻</td>
<td>Oxide ions (O₂⁻) lose electrons to form oxygen (O₂)</td>
<td>Aluminium ions (Al³⁺) gain electrons to form aluminium (Al)</td>
</tr>
<tr>
<td></td>
<td></td>
<td>2O₂⁻ − 4e⁻ → O₂</td>
<td>Al³⁺ + 3e⁻ → Al</td>
</tr>
</tbody>
</table>
Many metals are extracted from metal compounds in ores by heating with carbon in a reduction reaction. However, some metals cannot be extracted this way. This is because

- some metals are more reactive than carbon and/or
- some metals would react with carbon in the process.

Electrolysis is usually used to extract metals that cannot be extracted by heating with carbon. However, metals produced this way are expensive because of

- the high cost of thermal energy to melt the metal compounds and
- the high cost of the electricity for the process.

**Extraction of aluminium**

Aluminium is the second-most commonly used metal after iron/steel. It is too reactive to be extracted by heating with carbon and so is extracted by electrolysis.

The main ore of aluminium is bauxite which contains aluminium oxide. This has a very high melting point of 2072°C. The cost of the thermal...
Electrolysis of aqueous ionic compounds

When an ionic compound is dissolved in water, the products of electrolysis are often different to those when the compound is molten. In water, a small fraction of the molecules break down into hydrogen ions (H\(^+\)) and hydroxide ions (OH\(^-\)). These ions can be discharged instead of the ions in the ionic compound. More water molecules can break down if these ions are used up.

At each electrode there are two ions that could discharge, one from the ionic compound and one from the water. The one that is easier to discharge is the one that is discharged. Table 12.19 shows which ions are discharged at each electrode when inert electrodes are used. Inert electrodes are electrodes that will allow the electrolysis to take place but do not react themselves. Graphite electrodes are the most common inert electrodes used.

**Figure 12.28** The extraction of aluminium by electrolysis.

**TIP**

Remember that aluminium forms at the cathode and oxygen at the anode.

**Show you can...**

Complete the table below for the electrolysis of molten aluminium oxide.

**Table 12.18**

<table>
<thead>
<tr>
<th>Anode</th>
<th>Cathode</th>
</tr>
</thead>
<tbody>
<tr>
<td>Product</td>
<td></td>
</tr>
<tr>
<td>Half equation</td>
<td></td>
</tr>
<tr>
<td>Oxidation or reduction</td>
<td></td>
</tr>
<tr>
<td>Conditions for the electrolysis</td>
<td></td>
</tr>
</tbody>
</table>

**Electrolysis of aqueous ionic compounds**

Energy to melt aluminium oxide for electrolysis is very high. However, if the aluminium oxide is mixed with a substance called cryolite, the mixture melts at about 950°C and so the cost of thermal energy to melt this mixture is lower.

The electrodes for the process are made of graphite, a form of carbon (Figure 12.28). Aluminium ions (Al\(^{3+}\)) are attracted to the negative electrode where they gain electrons and form aluminium metal. As it is so hot, this is produced as a liquid and is run off at the bottom. Oxide ions (O\(^{2-}\)) are attracted to the positive electrode where they lose electrons and form oxygen. This oxygen reacts with the carbon anode and so the anode burns to produce carbon dioxide. This means that the anode has to be replaced regularly.

**Tip**

Cryolite is added to the aluminium oxide to reduce the cost of thermal energy.
Table 12.19 Products from the electrolysis of aqueous solutions of ionic compounds.

<table>
<thead>
<tr>
<th>Electrode</th>
<th>Positive electrode</th>
<th>Negative electrode</th>
</tr>
</thead>
<tbody>
<tr>
<td>Which ions are discharged</td>
<td>Oxygen is produced from the discharge of hydroxide ions (unless the ionic compound contains halide ions when these are discharged producing a halogen)</td>
<td>Positive ions discharged: Hydrogen is produced from the discharge of hydrogen ions, unless the ionic compound contains metal ions from a metal that is less reactive than hydrogen when the metal ions are discharged producing the metal.</td>
</tr>
</tbody>
</table>

Table 12.20 gives some examples to show which ions are discharged in the electrolysis of some aqueous solutions using inert electrodes.

**Table 12.20 Ion discharge in electrolysis.**

<table>
<thead>
<tr>
<th>Ionic compound (aqueous)</th>
<th>Sodium chloride</th>
<th>Copper(II) bromide</th>
<th>Silver nitrate</th>
<th>Potassium sulfate</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium chloride (NaCl(aq))</td>
<td>Na+ and H+</td>
<td>Cu2+ and H+</td>
<td>Ag+ and H+</td>
<td>K+ and H+</td>
</tr>
<tr>
<td>Copper(II) bromide (CuBr2(aq))</td>
<td>Cu2+</td>
<td>Ag+</td>
<td>H+</td>
<td></td>
</tr>
<tr>
<td>Silver nitrate (AgNO3(aq))</td>
<td></td>
<td></td>
<td>Ag+ and e−</td>
<td></td>
</tr>
<tr>
<td>Potassium sulfate (K2SO4(aq))</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Negative electrode**

<table>
<thead>
<tr>
<th>Product</th>
<th>Notes</th>
<th>Half equation</th>
<th>Process</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen, H2</td>
<td>Sodium is more reactive than hydrogen</td>
<td>2H+ + 2e− → H2</td>
<td>Reduction</td>
</tr>
<tr>
<td>Copper, Cu</td>
<td>Copper is less reactive than hydrogen</td>
<td>Cu2+ + 2e− → Cu</td>
<td>Reduction</td>
</tr>
<tr>
<td>Silver, Ag</td>
<td>Silver is less reactive than hydrogen</td>
<td>Ag+ + e− → Ag</td>
<td>Reduction</td>
</tr>
<tr>
<td>Hydrogen, H2</td>
<td>Potassium is more reactive than hydrogen</td>
<td>2H+ + 2e− → H2</td>
<td>Reduction</td>
</tr>
</tbody>
</table>

**Positive electrode**

<table>
<thead>
<tr>
<th>Product</th>
<th>Notes</th>
<th>Half equation</th>
<th>Process</th>
</tr>
</thead>
<tbody>
<tr>
<td>Chlorine, Cl2</td>
<td>Cl− is a halide ion</td>
<td>2Cl− − 2e− → Cl2</td>
<td>Oxidation</td>
</tr>
<tr>
<td>Bromine, Br2</td>
<td>Br− is a halide ion</td>
<td>2 Br− − 2e− → Br2</td>
<td>Oxidation</td>
</tr>
<tr>
<td>Oxygen, O2</td>
<td>NO3− is not a halide ion</td>
<td>4OH− − 4e− → 2H2O + O2</td>
<td>Oxidation</td>
</tr>
<tr>
<td>Oxygen, O2</td>
<td>SO4−2 is not a halide ion</td>
<td>4OH− − 4e− → 2H2O + O2</td>
<td>Oxidation</td>
</tr>
</tbody>
</table>

**Test yourself**

37 Copy and complete the table to show the products of electrolysis of some aqueous solutions of ionic compounds with inert electrodes.

**Table 12.21**

<table>
<thead>
<tr>
<th>Ionic compound (aqueous)</th>
<th>Product at the negative electrode (cathode)</th>
<th>Product at the positive electrode (anode)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Potassium iodide (KI)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Copper(II) chloride (CuCl2)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Magnesium sulfate (MgSO4)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Copper(II) nitrate (Cu(NO3)2)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Zinc bromide (ZnBr2)</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

38 For the electrolysis of aqueous sodium bromide solution using inert electrodes, write a half equation for the process at the electrode shown and state whether it is an oxidation or reduction process:

a) at the positive electrode
b) at the negative electrode.

39 For the electrolysis of aqueous copper(II) sulfate solution using inert electrodes, write a half equation for the process at the electrode shown and state whether it is an oxidation or reduction process:

a) at the positive electrode
b) at the negative electrode.
Investigate what happens when aqueous solutions are electrolysed using inert electrodes

A student made a hypothesis which stated, ‘Oxygen is produced at the positive electrode in the electrolysis of an aqueous solution, unless the compound contains halide ions.’ To test this hypothesis the apparatus was set up as shown. The electricity supply was switched on and the observations at each electrode recorded. The experiment was repeated using different aqueous solutions as electrolyte and the results obtained recorded in the table.

![Figure 12.29](image)

**Table 12.22**

<table>
<thead>
<tr>
<th>Electrode Type</th>
<th>Electrolyte</th>
<th>Observations</th>
<th>Test used for product</th>
<th>Identity of product</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cathode (negative electrode)</td>
<td>Copper(II) chloride</td>
<td>Red brown solid</td>
<td>Appearance</td>
<td>Red brown solid</td>
</tr>
<tr>
<td></td>
<td>Calcium nitrate</td>
<td>Colourless gas</td>
<td>Squeaky pop when a burning splint is inserted</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Silver sulfate</td>
<td>Grey solid</td>
<td>Appearance</td>
<td>Grey solid</td>
</tr>
<tr>
<td></td>
<td>Potassium bromide</td>
<td>Colourless gas</td>
<td>Squeaky pop when a burning splint is inserted</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Sodium iodide</td>
<td>Colourless gas</td>
<td>Squeaky pop when a burning splint is inserted</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Sulfuric acid</td>
<td>Colourless gas</td>
<td>Squeaky pop when a burning splint is inserted</td>
<td></td>
</tr>
<tr>
<td>Anode (positive electrode)</td>
<td>Copper(II) chloride</td>
<td>Colourless gas</td>
<td>Bleaches damp litmus paper</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Calcium nitrate</td>
<td>Colourless gas</td>
<td>Relights a glowing splint</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Silver sulfate</td>
<td>Colourless gas</td>
<td>Relights a glowing splint</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Potassium bromide</td>
<td>Yellow-orange solution</td>
<td>Be...</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Sodium iodide</td>
<td>Brown solution</td>
<td>Be...</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Sulfuric acid</td>
<td>Colourless gas</td>
<td>Be...</td>
<td></td>
</tr>
</tbody>
</table>

**Questions**

1. Why are graphite electrodes used?
2. Why are the aqueous solutions made up in distilled water?
3. What would you add to the circuit to show that a current is flowing?
4. Suggest why the experiment was only carried out long enough to make the necessary observations.
5. Write down the identity of the missing products at the cathode to complete the table.
6. Summarise the rules for deciding what is formed at the cathode when aqueous solutions of ionic compounds are electrolysed using inert electrodes.
7. Write down the identity of the missing products at the anode, to complete the table.
8. Using the results of this experiment, state if the hypothesis made by the student was correct.
Chapter review questions

1. The table shows the pH of some solutions.

<table>
<thead>
<tr>
<th>Solution</th>
<th>A</th>
<th>B</th>
<th>C</th>
<th>D</th>
<th>E</th>
</tr>
</thead>
<tbody>
<tr>
<td>pH</td>
<td>12</td>
<td>9</td>
<td>1</td>
<td>6</td>
<td>7</td>
</tr>
</tbody>
</table>

a) Which solution is the most acidic?
b) Which solution is the most alkaline?
c) Which solution is neutral?

2. Complete the following word equations, or write no reaction if no reaction would take place.

a) iron + oxygen
b) zinc + sulfuric acid
c) magnesium oxide + hydrochloric acid
d) sodium carbonate + nitric acid
e) potassium hydroxide + sulfuric acid
f) gold + magnesium nitrate
g) iron + copper sulfate
h) zinc + iron nitrate

3. Metals are extracted from compounds found in ores. Which method is used for each of the following metals?

a) magnesium
b) aluminium
c) copper
d) zinc
e) gold

4. Balance the following equations.

a) $Ca + O_2 \rightarrow CaO$
b) $NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$
c) $Fe_2O_3 + C \rightarrow Fe + CO$

5. Sodium hydroxide reacts with nitric acid to form the salt sodium nitrate.

a) Write a word equation for this reaction.
b) Write a balanced equation for this reaction.
c) Write an ionic equation for this reaction.
d) What type of reaction is this?

6. The salt magnesium sulfate is found in Epsom salts, which is used as a cure for constipation. It can be made by heating an excess of magnesium oxide with sulfuric acid. After the reaction is over, the left over magnesium oxide is filtered off. Crystallisation is used to produce the magnesium sulfate from the solution.

a) Why is an excess of magnesium oxide used?
b) What is the purpose of the filtration in this method?
c) Describe how the crystallisation of magnesium sulfate could be carried out.
d) Write a word equation for this reaction.
e) Write a balanced equation for this reaction.
7 Copy and complete the following table to show the products of electrolysis of some ionic compounds.

Table 12.24

<table>
<thead>
<tr>
<th>Ionic compound</th>
<th>Product at the negative electrode (cathode)</th>
<th>Product at the positive electrode (anode)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Molten magnesium bromide (MgBr₂)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Aqueous magnesium bromide (MgBr₂)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Molten potassium oxide (K₂O)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Molten sodium iodide (NaI)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Aqueous calcium nitrate (Ca(NO₃)₂)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Aqueous copper(II) chloride (CuCl₂)</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

8 Zn metal is produced when zinc oxide reacts with aluminium in a displacement reaction.

\[ 3\text{ZnO} + 2\text{Al} \rightarrow 3\text{Zn} + \text{Al}_2\text{O}_3 \]

a) Explain why aluminium displaces zinc.
b) Which substance is oxidised in this reaction? Give a reason for your answer.
c) Which substance is reduced in this reaction? Give a reason for your answer.

9 a) In the electrolysis of molten sodium chloride:
   i) Identify the products at the electrodes.
   ii) Write a half equation for the process at the negative electrode and state whether it is an oxidation or reduction process.
   iii) Write a half equation for the process at the positive electrode and state whether it is an oxidation or reduction process.
b) In the electrolysis of aqueous sodium chloride:
   i) Identify the products at the electrodes.
   ii) Write a half equation for the process at the negative electrode and state whether it is an oxidation or reduction process.
   iii) Write a half equation for the process at the positive electrode and state whether it is an oxidation or reduction process.
c) Explain why the electrolysis of molten and aqueous sodium chloride do not produce the same products.

10 Silver is produced when iron metal is added to a solution of silver nitrate.

\[ \text{Fe} + 2\text{AgNO}_3 \rightarrow \text{Fe(NO}_3)_2 + 2\text{Ag} \]
a) Write an ionic equation for this reaction.
b) Write two half equations for this reaction.
c) Identify the substance that is oxidised in this reaction and explain your answer.
d) Explain why this is a redox reaction.

11 Write a balanced equation for each of the following reactions.
   a) potassium + water
   b) copper carbonate + hydrochloric acid
   c) magnesium oxide + sulphuric acid
   d) zinc + silver nitrate
   e) aluminium + oxygen

12 10 cm³ of a solution of a strong acid had a pH 2. Some water was added to this to make the volume up to 1 dm³. What is the pH of the new solution? Explain your answer.
**Practice questions**

1. Which one of the following substances does not react with dilute hydrochloric acid at room temperature? [1 mark]
   - A) calcium carbonate
   - B) copper
   - C) magnesium
   - D) potassium hydroxide

2. The electrical conductivity of an electrolyte in electrolysis is due to the movement of particles through the substance when it is molten or in solution. Which of the following are the particles that move in the electrolyte? [1 mark]
   - A) atoms
   - B) electrons
   - C) ions
   - D) protons

3. Strips of four different metals were placed in solution of the nitrates of the same metals. Any reactions which occur are represented by a ✓ in the table. Which of the following is the correct order of reactivity with the most reactive first? [1 mark]
   ![Table 12.25]
   - A) P Q R S
   - B) Q R S P
   - C) R S P Q
   - D) Q R P S

4. Which one of the following could neutralise a solution of pH 10? [1 mark]
   - A) ammonia solution
   - B) hydrochloric acid
   - C) sodium hydroxide solution
   - D) water

5. Which one of the following pairs of substances could be used for the preparation of magnesium sulfate crystals? [1 mark]
   - A) dilute hydrochloric acid and magnesium carbonate
   - B) dilute nitric acid and magnesium oxide
   - C) dilute sulfuric acid and magnesium chloride
   - D) dilute sulfuric acid and magnesium carbonate.

6. What is the product formed at the anode in the electrolysis of aqueous sodium chloride solution? [1 mark]
   - A) chlorine
   - B) hydrogen
   - C) oxygen
   - D) sodium

7. Neutralisation occurs when an acid and an alkali react to form a salt and water.
   a) i) Copy and complete the table below to give the names and formulae of the ions present in all acids and alkalis. [4 marks]
   ![Table 12.26]
   - ii) Write an ionic equation for neutralisation, including state symbols. [2 marks]
   - b) Sulfuric acid solution can be neutralised using an alkali such as sodium hydroxide or adding a solid oxide such as copper(II) oxide.
     - i) Write a balanced equation for the reaction between sodium hydroxide and sulfuric acid. [2 marks]
     - ii) Write a balanced equation for the reaction between copper(II) oxide and sulfuric acid. [2 marks]
   - c) In the preparation of hydrated copper sulfate(II) crystals, an excess of copper(II) oxide was added to warm dilute sulfuric acid. The excess copper(II) oxide was removed by filtration. Describe how you would obtain pure dry crystals of hydrated copper(II) sulfate from the filtrate collected. [4 marks]

8. The description that follows is about a metal, other than zinc, which belongs to the reactivity series.
   - At room temperature the metal is a silver coloured solid. It reacts rapidly with cold dilute sulfuric acid to produce hydrogen. It conducts electricity. On heating in air the metal burns with a very bright flame leaving a white powder.
   - a) State one property from the description above that is common to all metals. [1 mark]
   - b) Give two reasons why the metal is NOT copper. [2 marks]
   - c) What part of the air reacts with the metal? [1 mark]
   - d) Suggest one metal which fits the description. [1 mark]
   - e) Write a word equation for the reaction of your chosen metal with sulfuric acid. [1 mark]
9 Hydrochloric acid can react with calcium hydroxide and with calcium. **Compare and contrast** the reaction of hydrochloric acid with calcium hydroxide with the reaction of hydrochloric acid with calcium. In your answer you must include the names of all products for each reaction and the observations for each reaction. [6 marks]

10 Aluminium is the most abundant metal in the Earth’s crust. Aluminium ore is first purified to give aluminium oxide and the metal is then extracted from the aluminium oxide by electrolysis.

a) What is meant by the term electrolysis?

b) Name the ore from which aluminium is extracted.

c) The electrolysis of the purified ore is carried out in the Hall-Héroult cell. The diagram below shows the cell used.

   ![Figure 12.30](image)

   ▲ Figure 12.30

   i) Name X and Z. [2 marks]

   ii) Y is the electrolyte. Name the substances in the electrolyte. [2 marks]

   iii) Why is the electrolysis cell kept at about 950°C? [2 marks]

   iv) Name the products formed at the positive and negative electrodes. [2 marks]

   v) Write half equations for the reactions taking place at each electrode. [2 marks]

   vi) Which electrode must be replaced regularly? Write a balanced symbol equation to explain your answer. [2 marks]

   d) Give a reason in terms of electrons why the extraction of aluminium in this process is a reduction reaction. [1 mark]

11 Some substances, for example molten lead bromide and aqueous sodium chloride, are described as electrolytes. Other substances for example copper metal, are conductors.

An experiment, to investigate the electrolysis of the electrolyte molten lead bromide, was set up as shown in the diagram below.

   ![Figure 12.31](image)

   ▲ Figure 12.31

a) Some pieces of apparatus in the diagram are labelled A–C. State the correct name for each piece of apparatus. [3 marks]

b) Name a piece of apparatus which could be connected in the circuit to show that an electric current is flowing through the molten lead bromide. [1 mark]

c) Copy and complete the table to state the names of the products, and the half equations for the electrodes. [4 marks]

<table>
<thead>
<tr>
<th>Electrode</th>
<th>Name of product</th>
<th>Half equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td></td>
<td></td>
</tr>
<tr>
<td>B</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

d) Why does this electrolysis need to be carried out in a fume cupboard? [1 mark]

e) Describe the differences in conduction of electricity by copper metal and molten lead bromide. [2 marks]

f) Name the product at each electrode when aqueous sodium chloride is electrolysed. [2 marks]
12.1 Chemical changes

Working scientifically: Measurements and uncertainties

We often use apparatus to make measurements in chemistry. For example, we often measure the mass, volume or temperature of a substance.

Measurement instruments which you should be able to use correctly in chemistry include:
- measuring cylinder (Figure 12.32)
- pipette (Figure 12.33)
- burette
- thermometer
- balance.

Measuring cylinders, pipettes and burettes are used to measure out volumes of liquids. Pipettes and burettes, which are used in titrations, measure volumes more accurately than measuring cylinders.

A meniscus is the curve seen at the top of a liquid in response to its container. When reading the volume of the liquid, the measurement at the bottom of the meniscus curve is read. This should be read at eye level (Figure 12.34).

The resolution of a piece of apparatus is the smallest change it can measure. For example, in Figure 12.34, the resolution of the burette is ±0.1 cm³.

When a measurement is made, there is always some doubt or uncertainty about its value. Uncertainty is often recorded after a measurement as a ± value.

The uncertainty can be estimated from the range of results that are obtained when an experiment is repeated several times.

**KEY TERMS**

Meniscus The curve at the surface of a liquid in a container.

Resolution The smallest change a piece of apparatus can measure.

Uncertainty The range of measurements within which the true value can be expected to lie.
For example, the volume of a gas produced in a reaction was measured five times. The results are 82, 77, 78, 96 and 80 cm³.

The mean value is found after excluding any anomalous results (96 cm³ is anomalous here as it is significantly different from all the others):

### Mean volume

\[
\text{Mean volume} = \frac{82 + 77 + 78 + 80}{4} = 79 \pm 3 \text{ cm}^3
\]

The mean is quoted to the nearest unit as all the values are measured to the nearest unit. The uncertainty is ±3 cm³ as the highest and lowest values are within 3 cm³ of the mean.

**Questions**

1. Record the volume of liquid in each measuring cylinder A-D. All scales are shown in cm³.

   ![Figure 12.35](image)

2. Record the volume of liquid in each of the burettes E-G shown below. All scales are shown in cm³.

   ![Figure 12.36](image)

3. Record the temperatures shown on each thermometer H-M. All scales are shown in ºC.

   ![Figure 12.37](image)

4. In each pair of diagrams, find the difference in mass between the first reading and the second reading. All scales are shown in grams.

   ![Figure 12.38](image)

5. Find the mean value and its uncertainty for the boiling point of a substance which was measured several times. Values measured were 124, 126, 123, 125 and 123ºC.

6. Find the mean value and its uncertainty for the mass of gas produced in a reaction which was measured several times. Values measured were 0.36, 0.18, 0.33 and 0.40 g.

7. Find the mean value and its uncertainty for the volume of acid needed to neutralise an alkali which was measured several times. Values measured were 25, 27, 26, 20 and 25 cm³.
Energy changes

Many people have a gas fire to keep warm at home. A chemical reaction takes place when the gas burns. This chemical reaction releases a lot of thermal energy that keeps us warm. This chapter looks at why some chemical reactions release thermal energy and increase the temperature while other reactions remove thermal energy and lower the temperature.

This chapter covers specification points 5.5.1.1 to 5.5.1.3 and is called Energy changes. It covers exothermic and endothermic reactions.
Exothermic and endothermic reactions

Energy changes in reactions
When chemical reactions take place, thermal energy is transferred to or from the surroundings. Some reactions make their surroundings hotter and some make them colder.

Exothermic reactions
In exothermic reactions, thermal energy is transferred from the chemicals to the surroundings. As there is more thermal energy, the temperature increases and so it gets hotter. Most chemical reactions are exothermic.

In some exothermic reactions, only a small amount of thermal energy is transferred and the temperature may only rise by a few degrees. In some reactions a lot of thermal energy is transferred and the surroundings get very hot. Sometimes there is so much thermal energy transferred that the reactants catch fire (Table 13.1).

Applications of exothermic reactions
Hand warmers are used by people in cold places to keep their hands warm (Figure 13.2). There are many different types, but they all work by using an exothermic reaction that transfers thermal energy to the surroundings.

Previously you could have learnt:
- Energy cannot be made or destroyed – it can only be transferred from one form to another (this is the law of conservation of energy).
- An energy change is a sign that a chemical reaction has taken place.
- Some chemical reactions lead to an increase in temperature and some to a decrease in temperature.

Test yourself on prior knowledge
1. Chemical energy and thermal energy are two forms of energy. Write down three forms of energy besides these two.
2. What is the law of conservation of energy?
### Table 13.1 Examples of exothermic reactions.

<table>
<thead>
<tr>
<th>Example</th>
<th>Comments</th>
</tr>
</thead>
</table>
| Oxidation reactions e.g. respiration (Figure 13.1) | Oxidation reactions take place when substances react with oxygen. Many oxidation reactions are exothermic. Respiration takes place in the cells of living creatures and is the reaction of glucose with oxygen. Thermal energy is released in this reaction. For example:  
\[
\text{glucose} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}  
\]
\[
C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O  
\]
| Combustion reactions e.g. burning fuels | Combustion reactions take place when substances react with oxygen and catch fire. They are a type of oxidation reaction. Fuels burning, such as methane (CH₄) in natural gas, are combustion reactions. These reactions are very exothermic, release a lot of thermal energy and catch fire. For example:  
\[
\text{methane} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}  
\]
\[
\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}  
\]
| Neutralisation reactions e.g. acids reacting with alkalis | Neutralisation reactions take place when acids react with bases. Many neutralisation reactions are exothermic. For example, when hydrochloric acid reacts with sodium hydroxide some thermal energy is transferred to the surroundings and the temperature rises by a few degrees. For example:  
\[
\text{hydrochloric acid} + \text{sodium hydroxide} \rightarrow \text{sodium chloride} + \text{water}  
\]
\[
\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}  
\]  

Self-heating cans can also be very useful (Figure 13.3). They can be used to provide hot drinks, such as coffee and hot chocolate, or even foods. Inside these cans, the food or drink is separated from a layer containing the chemicals used for heating. This is often calcium oxide and a bag of water. When a button is pressed by the user, the bag of water is punctured and the water mixes with the calcium oxide. The calcium oxide reacts with the water in an exothermic reaction. The thermal energy released heats up the food or drink.

### Endothermic reactions

In **endothermic reactions**, **thermal energy** is transferred from the **surroundings** to the **chemicals**. As there is less thermal energy, the temperature decreases and so it gets colder (Table 13.2). Endothermic reactions are less common than exothermic reactions.

#### Applications of endothermic reactions

Some sports injury packs act as a cold pack to put on an injury to prevent swelling (Figure 13.4). Inside the pack is a bag of water and a substance such as ammonium nitrate. When the pack is squeezed the water bag bursts and the ammonium nitrate dissolves in the water in an endothermic process.

### Table 13.2 Examples of endothermic reactions.

<table>
<thead>
<tr>
<th>Example</th>
<th>Comments</th>
</tr>
</thead>
</table>
| Decomposition, e.g. thermal decomposition of metal carbonates | Green copper carbonate decomposes into black copper oxide when heated  
\[
\text{CuCO}_3 \rightarrow \text{CuO} + \text{CO}_2  
\]  

\| Reaction of acids with metal hydrogen carbonates | Citric acid and sodium hydrogen carbonate are in sherbet sweets and react in your mouth in an endothermic reaction  
\[
\text{citric acid} + \text{sodium hydrogen carbonate} \rightarrow \text{sodium citrate} + \text{carbon dioxide} + \text{water}  
\]  

**TIP**

Remember that endothermic reactions get colder.
Exothermic and endothermic reactions

Test yourself

1. Is each of the following reactions endothermic or exothermic?
   a) the temperature started at 21°C and finished at 46°C
   b) the temperature started at 18°C and finished at 14°C
   c) the temperature started at 19°C and finished at 25°C

2. Is each of the following reactions endothermic or exothermic?
   a) burning alcohol
   b) thermal decomposition of iron carbonate
   c) reaction of vinegar (containing ethanoic acid) with baking powder (sodium hydrogencarbonate)
   d) reaction of vinegar (containing ethanoic acid) with sodium hydroxide

3. a) Why does the temperature increase when an exothermic reaction takes place in solution?
   b) Why does the temperature decrease when an endothermic reaction takes place in solution?

Show you can...

The reactions P and Q can be classified in different ways.

P: calcium carbonate → calcium oxide + carbon dioxide
Q: sodium hydroxide + nitric acid → sodium nitrate + water

Complete the table by placing one or more ticks (✔) in each row for reactions P and Q to indicate the terms which apply to each reaction. More than one term can apply to each reaction.

<table>
<thead>
<tr>
<th>P</th>
<th>Q</th>
</tr>
</thead>
<tbody>
<tr>
<td>Combustion</td>
<td>✔</td>
</tr>
<tr>
<td>Decomposition</td>
<td>✔</td>
</tr>
<tr>
<td>Neutralisation</td>
<td>✔</td>
</tr>
<tr>
<td>Oxidation</td>
<td>✔</td>
</tr>
<tr>
<td>Respiration</td>
<td>✔</td>
</tr>
<tr>
<td>Exothermic</td>
<td>✔</td>
</tr>
<tr>
<td>Endothermic</td>
<td>✔</td>
</tr>
</tbody>
</table>

Reaction profiles

Chemical reactions can only occur when particles collide with each other with enough energy to react. The minimum energy particles must have to react is called the activation energy.

We can show the relative energy of reactants and products in a reaction profile (Figure 13.5). This can also show the activation energy.

Figure 13.4 Sports injury packs use an endothermic reaction to keep the injury cold.
Investigate the variables that affect temperature change in reacting solutions – the temperature change in a neutralisation reaction

In an experiment the following method was followed:
1. 25 cm$^3$ of 40 g/dm$^3$ sodium hydroxide was measured out and placed in a polystyrene cup.
2. A burette was filled with 36.5 g/dm$^3$ hydrochloric acid.
3. The temperature of the sodium hydroxide was measured and recorded.
4. 5 cm$^3$ of hydrochloric acid was added from the burette to the plastic beaker, and the temperature recorded. Additional hydrochloric acid was added 5 cm$^3$ at a time and the temperature recorded until the total volume of hydrochloric acid added was 40 cm$^3$.
5. The experiment was repeated.
6. The results are shown in Table 13.4.

Table 13.4 The temperature recorded when increasing volumes of 36.5 g/dm$^3$ HCl was added to 25 cm$^3$ of 40 g/dm$^3$ NaOH.

<table>
<thead>
<tr>
<th>Volume of acid added in cm$^3$</th>
<th>Temperature in °C</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Experiment 1</td>
</tr>
<tr>
<td>0</td>
<td>19.4</td>
</tr>
<tr>
<td>5</td>
<td>21.6</td>
</tr>
<tr>
<td>10</td>
<td>24.0</td>
</tr>
<tr>
<td>15</td>
<td>25.2</td>
</tr>
<tr>
<td>20</td>
<td>25.4</td>
</tr>
<tr>
<td>25</td>
<td>26.2</td>
</tr>
<tr>
<td>30</td>
<td>25.4</td>
</tr>
<tr>
<td>35</td>
<td>25.2</td>
</tr>
<tr>
<td>40</td>
<td>25.1</td>
</tr>
</tbody>
</table>

Questions
1. In this experiment identify the:
   a) independent variable
   b) dependent variable
   c) key control variables
2. Why was the experiment repeated?
3. Why was a polystyrene cup used rather than a glass beaker?
4. Why should the solution in the polystyrene cup be stirred after each addition of acid?
5. Plot a graph of average temperature (y-axis) against volume of acid added (x-axis).
6. Describe the trend shown by the results plotted.
7. It is thought that the highest temperature is reached when complete neutralisation has occurred. Suggest how you would experimentally confirm that the highest temperature reached is the point at which neutralisation has occurred.
   A different experiment was carried out by adding 36.5 g/dm$^3$ solutions of different acids to 25 cm$^3$ of 40 g/dm$^3$ sodium hydroxide in a plastic beaker. The highest temperature reached was recorded and presented in Table 13.5.

Table 13.5 The highest temperature recorded when 40 cm$^3$ of different types of 1 mol/dm$^3$ acid of was added to 25 cm$^3$ of 1 mol/dm$^3$ NaOH.

<table>
<thead>
<tr>
<th>Type of acid</th>
<th>Hydrochloric acid</th>
<th>Sulfuric acid</th>
<th>Ethanoic acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>Highest temp in °C</td>
<td>26.2</td>
<td>26.5</td>
<td>25.2</td>
</tr>
<tr>
<td>Repeat highest temperature in °C</td>
<td>26.4</td>
<td>26.3</td>
<td>25.4</td>
</tr>
<tr>
<td>Average highest temperature in °C</td>
<td>26.3</td>
<td>26.4</td>
<td>25.3</td>
</tr>
</tbody>
</table>

Questions
8. In this experiment identify the:
   a) independent variable
   b) dependent variable
   c) key control variables
9. State two conclusions you can draw from the results.
**Example**

Draw a reaction profile diagram for the reaction below which is exothermic.

\[ \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \]

**Answer**

\[ \text{CO}_2 + 2\text{H}_2\text{O} \]

**Test yourself**

4. Look at the following reaction profiles.

   - **Reaction 1**
   - **Reaction 2**
   - **Reaction 3**

   a) Which reaction(s) is/are exothermic?
   b) Which reaction(s) is/are endothermic?

5. a) Sketch the reaction profile for the following reaction which is endothermic: \[ \text{CuCO}_3 \rightarrow \text{CuO} + \text{CO}_2 \]
   b) Draw an arrow to show the overall energy change for the reaction and label it **O**.
   c) Draw an arrow to show the activation energy for the reaction and label it **A**.

---

**Bond energies**

Breaking a chemical bond takes energy. For example, 436 kJ of energy is needed to break one mole of H—H covalent bonds. Due to the law of conservation of energy, 436 kJ of energy must be released when making one mole of H—H covalent bonds (Figure 13.9).

During a chemical reaction:

- **energy** must be supplied to **break bonds** in the reactants
- **energy** is released when **bonds** in the products are made.

The overall energy change for a reaction equals the difference between the energy needed to break the bonds in the reactants and the energy released when bonds are made in the products (Table 13.6).
Energy change = \text{energy needed breaking bonds in reactants} – \text{energy released making bonds in products}

Table 13.6

<table>
<thead>
<tr>
<th></th>
<th>Exothermic reaction</th>
<th>Endothermic reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Comparison of bond energies</td>
<td>More energy is released making new bonds than is needed to break bonds</td>
<td>More energy is needed to break bonds than is released making new bonds</td>
</tr>
<tr>
<td>Sign of energy change</td>
<td>–</td>
<td>+</td>
</tr>
</tbody>
</table>

Example

Find the energy change in the following reaction using the bond energies given.

Bond energies: \( \text{C–H} 412 \text{kJ} \), \( \text{O=O} 496 \text{kJ} \), \( \text{C=O} 743 \text{kJ} \), \( \text{O–H} 463 \text{kJ} \)

\[
\text{H} - \text{C} - \text{H} + 2 \text{O} = \text{O} \quad \text{O} = \text{C} = \text{O} + 2 \text{H} = \text{O}
\]

Explain why the reaction is exothermic or endothermic using bond energies.

**Answer**

Bonds broken: Bonds made:

\[
\begin{align*}
4\text{C–H} &= 4(412) \\
2\text{O=O} &= 2(496) \\
2\text{O–H} &= 2(463)
\end{align*}
\]

Total = 2640 kJ

Total = 3338 kJ

Energy change = \text{energy needed to break bonds} – \text{energy released making bonds}

\[
= 2640 – 3338 \\
= –698 \text{kJ}
\]

This reaction is exothermic because more energy is released making bonds than is needed to break bonds.

Example

Find the energy change in the following reaction using the bond energies given.

Bond energies: \( \text{H–H} 436 \text{kJ} \), \( \text{O=O} 496 \text{kJ} \), \( \text{O–H} 463 \text{kJ} \)

\[
2\text{H–H} + \text{O=O} \quad 2\text{H} = \text{O} \quad \text{O} = \text{C} = \text{O} + 2\text{H} = \text{O}
\]

Explain why the reaction is exothermic or endothermic using bond energies.

**Answer**

Bonds broken: Bonds made:

\[
\begin{align*}
2\text{H–H} &= 2(436) \\
\text{O=O} &= 496 \\
4\text{O–H} &= 4(463)
\end{align*}
\]

Total = 1368 kJ

Total = 1852 kJ

Energy change = \text{energy needed to break bonds} – \text{energy released making bonds}

\[
= 1368 – 1852 \\
= –484 \text{kJ}
\]

This reaction is exothermic because more energy is released making bonds than is needed to break bonds.
Exothermic and endothermic reactions

**Example**

Find the energy change in the following reaction using the bond energies given.

**Bond energies:**
- C–C: 348 kJ
- C–H: 412 kJ
- C–O: 360 kJ
- O–H: 463 kJ
- C–D: 612 kJ


\[
\begin{align*}
\text{HC} &\quad \text{H} \\
\text{H} &\quad \text{C} \quad \text{C} \\
\text{C} &\quad \text{H} \\
\text{O} &\quad \text{H}
\end{align*}
\]

**Answer**

**Bonds broken:**
- 5 C–H = 5(412) = 2060 kJ
- C–C = 348 kJ
- C–O = 360 kJ
- O–H = 463 kJ

**Total** = 3231 kJ

**Bonds made:**
- 4 C–H = 4(412) = 1648 kJ
- C–D = 612 kJ
- 2 O–H = 2(463) = 926 kJ

**Total** = 3186 kJ

Energy change = energy needed to break bonds
- energy released making bonds

= 3231 – 3186

= +45 kJ

This reaction is endothermic because more energy is needed breaking bonds than is released making bonds.

**Test yourself**

6 For each of the following reactions, calculate the energy change and explain why the reaction is endothermic or exothermic by discussing bond energies.

<table>
<thead>
<tr>
<th>Bond energy in kJ</th>
<th>C–C</th>
<th>C–H</th>
<th>C–Br</th>
<th>C–O</th>
<th>O–O</th>
<th>O–H</th>
<th>H–H</th>
<th>H–Br</th>
<th>N=N</th>
<th>N–H</th>
<th>Br–Br</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>348</td>
<td>612</td>
<td>412</td>
<td>276</td>
<td>743</td>
<td>496</td>
<td>463</td>
<td>436</td>
<td>366</td>
<td>944</td>
<td>388</td>
</tr>
</tbody>
</table>

**a)**

\[
\begin{align*}
\text{H} &\quad \text{H} \\
\text{Br} &\quad \text{Br} \\
\text{Br} &\quad \text{Br}
\end{align*}
\]

**b)**

\[
\begin{align*}
\text{N} &\quad \text{N} \\
\text{H} &\quad \text{H}
\end{align*}
\]

**c)**

\[
\begin{align*}
\text{H} &\quad \text{C} \quad \text{C} \\
\text{H} &\quad \text{H} \\
\text{C} &\quad \text{H} \\
\text{Br} &\quad \text{Br}
\end{align*}
\]

**d)**

\[
\begin{align*}
\text{H} &\quad \text{C} \quad \text{C} \quad \text{C} \\
\text{H} &\quad \text{H} \quad \text{H} \\
\text{O} &\quad \text{O} \\
\text{H} &\quad \text{H}
\end{align*}
\]

**Show you can...**

The energy change is –30 kJ in the following reaction:

\[
\text{I} + \text{I} + \text{Cl} = 2\text{I} + \text{Cl}
\]

**Bond energies:**
- I–I = 150 kJ
- Cl–Cl = 242 kJ

Calculate the bond energy of I–Cl.
Chapter review questions

1. The following reactions all took place in solution in a beaker. The temperature before and after the chemicals were mixed was recorded in each case. Decide whether each reaction is exothermic or endothermic.

Table 13.8

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Start temperature in °C</th>
<th>End temperature in °C</th>
</tr>
</thead>
<tbody>
<tr>
<td>Reaction 1</td>
<td>21</td>
<td>15</td>
</tr>
<tr>
<td>Reaction 2</td>
<td>20</td>
<td>27</td>
</tr>
<tr>
<td>Reaction 3</td>
<td>22</td>
<td>67</td>
</tr>
</tbody>
</table>

2. Copy and complete the spaces in the following sentences.
   In an exothermic reaction, thermal energy is transferred from the chemicals to their surroundings and so the temperature ________________. In an __________________ reaction, thermal energy is transferred away from the surroundings to the chemicals and so the temperature ________________.

3. Copy and complete the spaces in the following sentences.
   Chemical reactions can only take place when particles _____________ with each other and have enough energy. The minimum energy particles need to react is called the _____________ energy.

4. The reaction profile for a reaction is shown.

   ![Figure 13.11](image)

   ▲ Figure 13.11

   a) Give the letter of the arrow that shows the activation energy for the reaction.

   b) Give the letter of the arrow that shows the overall energy change for the reaction.

   c) Is this reaction endothermic or exothermic?

5. Decide whether each of the following reactions is likely to be endothermic or exothermic.

   a) burning magnesium

   b) decomposition of silver oxide

   c) reaction inside a sports injury cold pack

   d) reaction inside a self-heating food can

   e) neutralisation of sulfuric acid by sodium hydroxide

   f) neutralisation of sulfuric acid by sodium hydrogen carbonate
6 a) Calculate the energy change for the following reaction using these bond energies.

\[ \text{H—H} = 436 \text{ kJ}, \text{Cl—Cl} = 242 \text{ kJ}, \text{H—Cl} = 431 \text{ kJ} \]

\[ \text{H—H + Cl—Cl} \rightarrow 2 \text{H—Cl} \]

b) Is this reaction endothermic or exothermic? Explain your answer by discussing bond energies.

7 Calculate the energy change for the following reaction using these bond energies.

\[ \text{N—H} = 388 \text{ kJ}, \text{N—N} = 158 \text{ kJ}, \text{N≡N} = 944 \text{ kJ}, \text{O—H} = 463 \text{ kJ}, \text{O≡O} = 496 \text{ kJ} \]

\[ \text{H} \text{N} \text{N} + \text{O} \equiv \text{O} \rightarrow \text{N≡N} + 2 \text{H—O—H} \]
Practice questions

1 Which one of the following is an endothermic reaction? [1 mark]
   A the combustion of hydrogen
   B the reaction of citric acid and sodium hydrogencarbonate
   C the reaction of magnesium and hydrochloric acid
   D the reaction of sodium hydroxide and hydrochloric acid

2 When potassium hydroxide dissolves in water the temperature of the solution rises. Which of the following is this an example of?
   A an endothermic change
   B an exothermic change
   C a neutralisation reaction
   D a thermal decomposition reaction

3 a) For each of the reactions A to E, choose the appropriate word from the list below to describe the type of reaction. Each word may be used once, more than once or not at all. [5 marks]
   combustion  decomposition  neutralisation  oxidation  reduction
   A copper carbonate → copper oxide + carbon dioxide
   B ethanoic acid + sodium hydroxide → sodium ethanoate + water
   C magnesium + oxygen → magnesium oxide
   D methane + oxygen → carbon dioxide + water
   E sodium hydroxide + hydrochloric acid → sodium chloride + water
   b) Explain the difference between an exothermic reaction and an endothermic reaction. [2 marks]
   c) For each of the reactions A to E above decide if it is an exothermic or an endothermic reaction. [5 marks]
   d) Describe how you would experimentally prove that the reaction between magnesium and hydrochloric acid is an exothermic reaction. [4 marks]

4 Photosynthesis is an endothermic process used by plants to produce carbohydrates.
   \[ 6\text{CO}_2 + 6\text{H}_2\text{O} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2 \]
   a) What is meant by the term endothermic? [1 mark]
   b) Describe what is meant by the term activation energy. [1 mark]
   c) Draw a labelled reaction profile for this reaction. You must show: the position of the reactants and products; the activation energy; and the energy change of the reaction. [4 marks]

5 The diagram shows the profile for a reaction with a catalyst and without a catalyst.

   ![Diagram](Figure 13.12)

   a) Is this an exothermic or an endothermic reaction? [1 mark]
   b) Does a catalyst have an effect on the overall energy change of the reaction? [1 mark]
   c) On a copy of the diagram label the activation energy for the catalysed reaction as A and the activation energy for the uncatalysed reaction as B. [2 marks]
   d) From the information shown in the graph, state the effect of a catalyst on the activation energy of a reaction. [1 mark]

6 Hydrogen reacts with fluorine to form hydrogen fluoride.
   \[ \text{H}_2 + \text{F}_2 \rightarrow 2\text{HF} \]
   Use the bond energies in the table to calculate the energy change for this reaction. Explain if the reaction is exothermic or endothermic. [4 marks]

<table>
<thead>
<tr>
<th>Bond</th>
<th>Bond energy in kJ</th>
</tr>
</thead>
<tbody>
<tr>
<td>H–H</td>
<td>436</td>
</tr>
<tr>
<td>F–F</td>
<td>158</td>
</tr>
<tr>
<td>H–F</td>
<td>568</td>
</tr>
</tbody>
</table>

7 Ethanol burns in oxygen.
   \[ \text{H}_2\text{C} \text{O} + 3\text{O}_2 \rightarrow 2\text{CO} + 3\text{H}_2\text{O} \]
   Use the bond energies below to calculate the energy change for this reaction. [3 marks]

<table>
<thead>
<tr>
<th>Bond energies:</th>
</tr>
</thead>
<tbody>
<tr>
<td>C–H 412 kJ</td>
</tr>
<tr>
<td>O=O 496 kJ</td>
</tr>
<tr>
<td>C=O 743 kJ</td>
</tr>
<tr>
<td>O–H 463 kJ</td>
</tr>
<tr>
<td>C–C 348 kJ</td>
</tr>
<tr>
<td>C–O 360 kJ</td>
</tr>
</tbody>
</table>
Working scientifically: Identifying variables when planning experiments

When carrying out an experiment different variables are used. Variables are the things that can change. When we plan experiments, we choose to change some variables while keeping others the same.

A variable is a physical, chemical or biological quantity or characteristic that can have different values. It may be for example, temperature, mass, volume, pH or even the type of chemical used in an experiment.

There are different types of variables which you need to be familiar with.

A continuous variable has values that are numbers. Mass, temperature and volume are examples of continuous variables. The values of these variables can either be found by counting (e.g. the number of drops) or by measurement (e.g. the temperature).

A categoric variable is one which is best described by words. Variables such as the type of acid or the type of metal are categoric variables.

Scientists often plan experiments to investigate if there is a relationship between two variables, the independent and the dependent variable.

The independent variable is the variable for which values are changed or selected by the investigator (i.e. it is the one which you deliberately change during an experiment).

The dependent variable is one which may change as a result of changing the independent variable. It is the one which is measured for each and every change in the independent variable.

A control variable is one which may, in addition to the independent variable, affect the outcome of the investigation. Control variables must be kept constant during an experiment to make it a fair test.

Questions

1. Decide if the following are categoric or continuous variables.
   a) temperature
   b) type of acid
   c) volume
   d) name of gas produced
   e) concentration of solution
   f) volume of gas produced
   g) mass
   h) name of metal
   i) colour of solution
   j) drops of acid

Figure 13.13 The volume of carbon dioxide gas produced over time, from the reaction between calcium carbonate and 36.5 g/dm³ HCl was recorded using this apparatus. The experiment was repeated using different concentrations of HCl. Can you identify the independent, dependent and controlled variables?
As an example, the variables are shown below for an experiment to find the effect of temperature on the rate of the reaction between magnesium and excess hydrochloric acid. The rate was measured by timing how long it took for all the magnesium to react.

Table 13.10

<table>
<thead>
<tr>
<th>Independent variable</th>
<th>Dependent variable</th>
<th>Control variables</th>
</tr>
</thead>
<tbody>
<tr>
<td>Temperature</td>
<td>Time taken for all the magnesium to react</td>
<td>The mass of magnesium, The surface area of magnesium, The volume of hydrochloric acid, The concentration of hydrochloric acid</td>
</tr>
</tbody>
</table>

Questions

2 For the following experiments identify the
(i) independent variable
(ii) dependent variable
(iii) control variables.

a) Some magnesium was added to hydrochloric acid and the temperature recorded. The experiment was repeated several times using different volumes of hydrochloric acid.

b) In the reaction between copper carbonate and hydrochloric acid the time taken for a mass of copper carbonate to all be used up was recorded. The experiment was repeated using different masses of copper carbonate.

c) 2g of magnesium was added to copper sulfate solution and the highest temperature reached was recorded. The experiment was repeated using five different concentrations of copper sulfate.

d) In an experiment to find the effect of stirring on speed of dissolving, the time taken to dissolve some copper sulfate in water was measured. This was repeated stirring the solution.

e) The volume of carbon dioxide gas produced when calcium carbonate reacts with hydrochloric acid was measured, and the experiment repeated using different masses of calcium carbonate.

f) The temperature of nitric acid was recorded before and after some sodium hydroxide was added. The experiment was repeated using sulfuric acid, ethanoic acid and methanoic acid.
This chapter is called Formulae and Equations and brings together points from throughout the specification.

Chemists use formulae and equations as a quick way of identifying substances and showing what happens in chemical reactions. Being formula literate is vital for any chemist. This chapter looks at writing formulae and understanding what they mean, as well as how to write chemical, ionic and half equations.
Chemists use formulae a lot and it is important that you are formula literate meaning that you can write and recognise formulae.

**Elements**

The formula for most elements is just its symbol. For example, the formula of argon is Ar and that of magnesium is Mg (Table 14.1).

However, this is not the case for elements made of molecules. Many of these molecules contain two atoms (called diatomic molecules) such as hydrogen (H₂) and oxygen (O₂). Some elements that are made of molecules contain more than two atoms in their molecules, such as phosphorus molecules which contain four atoms (P₄).

### Table 14.1 Formulae of some common elements.

<table>
<thead>
<tr>
<th>Common elements whose formula is the symbol</th>
<th>Common elements whose formula is not the symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Al  aluminium</td>
<td>Br₂  bromine</td>
</tr>
<tr>
<td>Ar  argon</td>
<td>C₆₀  carbon (buckminsterfullerene)</td>
</tr>
<tr>
<td>Be  beryllium</td>
<td>Cl₂  chlorine</td>
</tr>
<tr>
<td>B   boron</td>
<td>F₂   fluorine</td>
</tr>
<tr>
<td>Ca  calcium</td>
<td>H₂   hydrogen</td>
</tr>
<tr>
<td>C   carbon (diamond)</td>
<td>I₂   iodine</td>
</tr>
<tr>
<td>C   carbon (graphite)</td>
<td>N₂   nitrogen</td>
</tr>
<tr>
<td>Cu  copper</td>
<td>O₂   oxygen</td>
</tr>
<tr>
<td>Au  gold</td>
<td>P₄   phosphorus</td>
</tr>
<tr>
<td>He  helium</td>
<td>Zn   zinc</td>
</tr>
<tr>
<td>Fe  iron</td>
<td></td>
</tr>
</tbody>
</table>

**KEY TERM**

*Diatomic molecule* A molecule containing two atoms.

**Compounds**

Some common compounds

It is very useful to know the formula of some common compounds. Some are listed in Table 14.2.

### Table 14.2 Formulae of some common compounds.

<table>
<thead>
<tr>
<th>Common compounds</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>NH₃</td>
<td></td>
</tr>
<tr>
<td>CO₂</td>
<td></td>
</tr>
<tr>
<td>CO</td>
<td></td>
</tr>
<tr>
<td>CH₄</td>
<td></td>
</tr>
<tr>
<td>NO</td>
<td></td>
</tr>
<tr>
<td>NO₂</td>
<td></td>
</tr>
<tr>
<td>SO₂</td>
<td></td>
</tr>
<tr>
<td>SO₃</td>
<td></td>
</tr>
<tr>
<td>H₂O</td>
<td></td>
</tr>
</tbody>
</table>
Ionic compounds

Compounds made from metals combined with non-metals have an ionic structure. The formula of each of these compounds can be worked out using ion charges. The charges of common ions are shown in Tables 14.3 and 14.4.

Table 14.3 Positive ions.

<table>
<thead>
<tr>
<th>Group 1 ions (form 1+ ions)</th>
<th>Group 2 ions (form 2+ ions)</th>
<th>Group 3 ions (form 3+ ions)</th>
<th>Others</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li+ lithium</td>
<td>Mg2+ magnesium</td>
<td>Al3+ aluminium</td>
<td></td>
</tr>
<tr>
<td>Na+ sodium</td>
<td>Ca2+ calcium</td>
<td></td>
<td></td>
</tr>
<tr>
<td>K+ potassium</td>
<td>Ba2+ barium</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Others</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Table 14.4 Negative ions.

<table>
<thead>
<tr>
<th>Group 6 ions (form 2- ions)</th>
<th>Group 7 ions (form 1- ions)</th>
<th>Others</th>
</tr>
</thead>
<tbody>
<tr>
<td>O2– oxide</td>
<td>F– fluoride</td>
<td></td>
</tr>
<tr>
<td>S2– sulfide</td>
<td>Cl– chloride</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Br– bromide</td>
<td></td>
</tr>
<tr>
<td></td>
<td>I– iodide</td>
<td></td>
</tr>
</tbody>
</table>
|                             |                             | CO3– carbonate
|                             |                             | OH– hydroxide
|                             |                             | NO3– nitrate
|                             |                             | SO4– sulfate

In an ionic substance the total number of positive charges must equal the total number of negative charges. This allows us to work out the formula of ionic substances.

Examples

Sodium oxide: contains sodium ions (Na+) and oxide ions (O2–)

There must be the same number of positive and negative charges, so we need two Na+ ions (total of two positive charges) for every one O2– ion (two negative charges)

Na+ O2–

Formula = Na2O

Iron(III) sulfide: contains iron(III) ions (Fe3+) and sulfide ions (S2–)

There must be the same number of positive and negative charges, so we need two Fe3+ ions (total of six positive charges) for every three S2– ions (six negative charges)

Fe3+ S2–

Formula = Fe2S3

Some ions contain atoms of different elements. Examples include sulfate (SO4–), hydroxide (OH–) and nitrate (NO3–). These are sometimes called compound ions or molecular ions. If you need to write more than one of these in a formula, then these ions should be placed in a bracket.

TIP
You should be able to work out the charge of ions of elements in Groups 1, 2, 6 and 7.

TIP
The charge on the nitrate (NO3–) ion is 1–, not 3–. The charge on the ammonium (NH4+) ion is 1+ not 4+.
Example
Magnesium hydroxide: contains magnesium ions (Mg$^{2+}$) and hydroxide ions (OH$^-$)

There must be the same number of positive and negative charges, so we need one Mg$^{2+}$ ion (total of two positive charges) for every two OH$^-$ ions (two negative charges)

\[
\text{Mg}^{2+} \quad \text{OH}^- \\
\text{OH}^- \\
\text{Formula} = \text{Mg(OH)}_2
\]

Table 14.5 Ions of some common compounds.

<table>
<thead>
<tr>
<th>Name</th>
<th>+ Ions</th>
<th>– Ions</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium chloride</td>
<td>Na$^+$ (1+ charge)</td>
<td>Cl$^-$ (1– charge)</td>
<td>NaCl</td>
</tr>
<tr>
<td>Magnesium chloride</td>
<td>Mg$^{2+}$ (2+ charges)</td>
<td>Cl$^-$ (2– charges)</td>
<td>MgCl$_2$</td>
</tr>
<tr>
<td>Magnesium sulfide</td>
<td>Mg$^{2+}$ (2+ charges)</td>
<td>S$^{2-}$ (2– charges)</td>
<td>MgS</td>
</tr>
<tr>
<td>Copper(II) sulfate</td>
<td>Cu$^{2+}$ (2+ charges)</td>
<td>SO$_4^{2-}$ (2– charges)</td>
<td>CuSO$_4$</td>
</tr>
<tr>
<td>Sodium carbonate</td>
<td>Na$^+$ (2+ charges)</td>
<td>CO$_3^{2-}$ (2– charges)</td>
<td>Na$_2$CO$_3$</td>
</tr>
<tr>
<td>Ammonium sulfate</td>
<td>NH$_4^+$ (2+ charges)</td>
<td>SO$_4^{2-}$ (2– charges)</td>
<td>(NH$_4$)$_2$SO$_4$</td>
</tr>
<tr>
<td>Calcium nitrate</td>
<td>Ca$^{2+}$ (2+ charges)</td>
<td>NO$_3^-$ (2– charge)</td>
<td>Ca(NO$_3$)$_2$</td>
</tr>
<tr>
<td>Aluminium oxide</td>
<td>Al$^{3+}$ (6+ charges)</td>
<td>O$_2^-$ (6– charges)</td>
<td>Al$_2$O$_3$</td>
</tr>
<tr>
<td>Iron(III) hydroxide</td>
<td>Fe$^{3+}$ (3+ charges)</td>
<td>OH$^-$ (3– charges)</td>
<td>Fe(OH)$_3$</td>
</tr>
</tbody>
</table>

Test yourself

1 Write the formula of each of the following elements and compounds.
   a) copper
   b) hydrogen
   c) carbon dioxide
   d) argon
   e) silver
   f) oxygen
   g) ammonia
   h) chlorine
   i) carbon (diamond)
   j) carbon (buckminsterfullerene)
   k) sulfur dioxide
   l) methane

2 Write the formula of each of the following ionic compounds.
   a) potassium oxide
   b) sodium sulfate
   c) aluminium fluoride
   d) iron(II) sulfide
   e) copper(II) nitrate
   f) lithium carbonate
   g) ammonium bromide
   h) barium hydroxide
   i) silver nitrate
   j) aluminium sulfate
   k) strontium oxide
   l) potassium selenide

3 Name each of the following substances.
   a) Br$_2$
   b) Na
   c) Cu
   d) CO
   e) SO$_3$
   f) CaO
   g) AlF$_3$
   h) CuS
   i) KNO$_3$
   j) (NH$_4$)$_2$CO$_3$
   k) FeO
   l) Fe$_2$O$_3$
Classifying substances

**Structure types**

It is very useful to be able to identify what type of structure a substance has from its name or formula. Table 14.6 gives some general guidance on this.

<table>
<thead>
<tr>
<th>Structure type</th>
<th>Description of structure</th>
<th>Which substances have this structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>Monatomic</td>
<td>Made of individual atoms</td>
<td>Group 0 elements</td>
</tr>
<tr>
<td>Molecular</td>
<td>Made of individual molecules</td>
<td>Some non-metal elements (e.g. H₂, C₆₀, N₂, O₂, F₂, P₄, Cl₂, Br₂, I₂)</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Compounds made from non-metals (e.g. CH₄, CO₂, H₂O, NH₃, C₆H₁₂O₆)</td>
</tr>
<tr>
<td>Giant covalent</td>
<td>Lattice of atoms joined by covalent bonds</td>
<td>Diamond (C), graphite (C), graphene (C), silicon (Si), silicon dioxide (SiO₂)</td>
</tr>
<tr>
<td>Ionic</td>
<td>Lattice of positive and negative ions</td>
<td>Compounds made from metals combined with non-metals (e.g. NaCl, Fe₂O₃, CuSO₄)</td>
</tr>
<tr>
<td>Metallic</td>
<td>Lattice of metal atoms in a cloud of delocalised outer shell electrons</td>
<td>Metals (e.g. Cu, Fe, Al, Na, Ca, Mg, Au, Ag, Pt)</td>
</tr>
</tbody>
</table>

**Acids, bases, alkalis and salts**

Some compounds act as acids, bases, alkalis or salts. It is very useful if you can identify an acid, base, alkali or salt although not all substances are one of these (Table 14.7).

<table>
<thead>
<tr>
<th>Acids</th>
<th>Substances that react with water to release H⁺ ions</th>
<th>Common acids:</th>
<th>H₂SO₄  sulfuric acid</th>
<th>HCl  hydrochloric acid</th>
<th>HNO₃  nitric acid</th>
<th>H₃PO₄  phosphoric acid</th>
<th>CH₃COOH  ethanoic acid</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Substances that react with acids to form a salt and water (and sometimes carbon dioxide as well)</td>
<td>Common bases:</td>
<td>Metal oxides e.g. CaO, Na₂O</td>
<td>Metal hydroxides e.g. Ca(OH)₂, NaOH</td>
<td>Metal carbonates e.g. CaCO₃, Na₂CO₃</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>Alkalis:</td>
<td>Substances that react with water to release OH⁻ ions</td>
<td>(they are a special type of water-soluble base)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>NH₃  ammonia</td>
<td>plus water-soluble metal hydroxides:</td>
<td>NaOH  sodium hydroxide</td>
<td>KOH  potassium hydroxide</td>
<td>Ca(OH)₂  calcium hydroxide</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>Common salts:</td>
<td>Sulfates from sulfuric acid</td>
<td>Chlorides from hydrochloric acid</td>
<td>Nitrates from nitric acid</td>
<td>Phosphates from phosphoric acid</td>
<td>Ethanoates from ethanoic acid</td>
</tr>
</tbody>
</table>

Alkalis are a special type of base and so any substance that is an alkali is also a base.

**Acid–base character of oxides**

Most metal oxides are basic (Table 14.8). For example, calcium oxide (CaO) is used as a base to neutralise acidic soil on farms.

Most non-metal oxides are acidic. For example, carbon dioxide (CO₂) dissolves in rain water to make rain naturally slightly acidic.
Table 14.8 Oxides.

<table>
<thead>
<tr>
<th>Type of oxide</th>
<th>Metal oxides</th>
<th>Non-metal oxides</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acidic or basic</td>
<td>Basic (react with acids)</td>
<td>Acidic (react with bases)</td>
</tr>
<tr>
<td>Examples</td>
<td>Calcium oxide (CaO)</td>
<td>Carbon dioxide (CO₂)</td>
</tr>
<tr>
<td></td>
<td>Sodium oxide (Na₂O)</td>
<td>Sulfur dioxide (SO₂)</td>
</tr>
<tr>
<td></td>
<td>Copper oxide (CuO)</td>
<td>Phosphorus oxide (P₄O₁₀)</td>
</tr>
</tbody>
</table>

Test yourself

4 What type of structure does each of the following substances have?
   a) lead (Pb)
   c) potassium iodide (KI)
   e) diamond (C)
   g) ethanol (C₂H₅OH)
   i) chromium (Cr)
   k) sulfur dioxide (SO₂)
   b) argon (Ar)
   d) oxygen (O₂)
   f) methane (CH₄)
   h) aluminium oxide (Al₂O₃)
   j) silicon dioxide (SiO₂)
   l) potassium nitrate (KNO₃)

5 Classify each of these substances as an acid, base, alkali or salt.
   a) Fe₂O₃
   b) Na₂SO₄
   c) KOH
   d) ZnCO₃
   e) HNO₃
   f) Ca(NO₃)₂
   g) NH₃
   h) K₂O
   i) HCl
   j) MgCl₂
   k) NaBr
   l) H₂SO₄

6 Classify each of the following oxides as acidic or basic.
   a) NO₂
   b) K₂O
   c) MgO
   d) SiO₂

Common reactions

There are some general reactions that are useful to know. Many of these involve acids and/or metals. Remember that hydrochloric acid produces chloride salts, sulfuric acid produces sulfate salts and nitric acid produces nitrate salts.

Examples

- element + oxygen → oxide of element
  e.g. calcium + oxygen → calcium oxide

- compound + oxygen → oxides of each element in compound
  e.g. methane (CH₄) + oxygen → carbon dioxide + water

- water + metal → metal hydroxide + hydrogen (for metals that react with water)
  e.g. water + sodium → sodium hydroxide + hydrogen

- acid + metal → salt + hydrogen (for metals that react with dilute acids)
  e.g. hydrochloric acid + magnesium → magnesium chloride + hydrogen

- acid + metal oxide → salt + water
  e.g. sulfuric acid + copper oxide → copper sulfate + water

- acid + metal hydroxide → salt + water
  e.g. nitric acid + potassium hydroxide → potassium nitrate + water

- acid + metal carbonate → salt + water + carbon dioxide
  e.g. hydrochloric acid + calcium carbonate → calcium chloride + water + carbon dioxide

- acid + ammonia → ammonium salt
  e.g. nitric acid + ammonia → ammonium nitrate
Chemical reactions

Table 14.9 Chemical reactions.

Chemical reactions take place in one of the three ways shown in Table 14.9.

<table>
<thead>
<tr>
<th>Way in which the reaction takes place</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 Transfer of electrons</td>
<td><strong>Metals reacting with non-metals</strong></td>
</tr>
<tr>
<td></td>
<td>e.g. sodium + chlorine → sodium chloride</td>
</tr>
<tr>
<td></td>
<td>2Na + Cl₂ → 2NaCl</td>
</tr>
<tr>
<td></td>
<td>Sodium atoms lose electrons to form sodium ions. These electrons are transferred to chlorine atoms which form chloride ions. This forms the ionic compound sodium chloride</td>
</tr>
<tr>
<td>2 Sharing of electrons</td>
<td><strong>Displacement reactions</strong></td>
</tr>
<tr>
<td></td>
<td>e.g. zinc + copper sulfate → zinc sulfate + copper</td>
</tr>
<tr>
<td></td>
<td>Zn + CuSO₄ → ZnSO₄ + Cu</td>
</tr>
<tr>
<td></td>
<td>Zinc atoms lose electrons to form zinc ions. These electrons are transferred to copper ions in copper sulfate forming copper atoms.</td>
</tr>
<tr>
<td>3 Transfer of protons</td>
<td><strong>Sharing of electrons</strong></td>
</tr>
<tr>
<td></td>
<td>Non-metals reacting with non-metals</td>
</tr>
<tr>
<td></td>
<td>e.g. hydrogen + oxygen → water</td>
</tr>
<tr>
<td></td>
<td>H₂ + O₂ → 2H₂O</td>
</tr>
<tr>
<td></td>
<td>Hydrogen atoms share electrons with oxygen atoms to form covalent bonds in water</td>
</tr>
<tr>
<td></td>
<td><strong>Acids reacting with alkalis</strong></td>
</tr>
<tr>
<td></td>
<td>e.g. hydrochloric acid + sodium hydroxide → sodium chloride + water</td>
</tr>
<tr>
<td></td>
<td>HCl + NaOH → NaCl + H₂O</td>
</tr>
<tr>
<td></td>
<td>H⁺ ions (protons) are transferred from the hydrochloric acid to the OH⁻ ions in sodium hydroxide to form water</td>
</tr>
<tr>
<td></td>
<td><strong>Acids reacting with metal oxides</strong></td>
</tr>
<tr>
<td></td>
<td>e.g. sulfuric acid + copper oxide → copper sulfate + water</td>
</tr>
<tr>
<td></td>
<td>H₂SO₄ + CuO → CuSO₄ + H₂O</td>
</tr>
<tr>
<td></td>
<td>H⁺ ions (protons) are transferred from the sulfuric acid to the O²⁻ ions in copper oxide to form water</td>
</tr>
</tbody>
</table>

Test yourself

7 Write a word equation for each of the following reactions with oxygen.
   a) magnesium (Mg) + oxygen
   b) hydrogen sulfide (H₂S) + oxygen
   c) phosphorus (P₄) + oxygen
   d) silane (SiH₄) + oxygen
   e) propane (C₃H₈) + oxygen
   f) methanol (CH₃OH) + oxygen
   g) magnesium hydroxide + sulfuric acid
   h) calcium + water
   i) copper carbonate + nitric acid
   j) ammonia + sulfuric acid
   k) magnesium oxide + nitric acid
   l) cobalt + hydrochloric acid

8 Write a word equation for each of the following reactions.
   a) potassium + water
   b) nitric acid + zinc
   c) sulfuric acid + nickel oxide
   d) hydrochloric acid + potassium hydroxide
   e) nitric acid + sodium carbonate
   f) hydrochloric acid + ammonia
   g) magnesium hydroxide + sulfuric acid
   h) calcium + water
   i) copper carbonate + nitric acid
   j) ammonia + sulfuric acid
   k) magnesium oxide + nitric acid
   l) cobalt + hydrochloric acid

9 Which of the following reactions involves the transfer of electrons?
   1. sharing of electrons?
   2. transfer of protons?
   a) nitric acid + sodium oxide → sodium nitrate + water
   b) aluminium + iron oxide → aluminium oxide + iron
   c) hydrogen + sulfur → hydrogen sulfide
   d) aluminium + bromine → aluminium bromide
   e) hydrochloric acid + sodium carbonate → sodium chloride + water + carbon dioxide
Balancing equations

Word equations show the names of the reactants and products in a reaction. A balanced symbol (formula) equation shows the formula of each substance and how many particles of each are involved in the reaction. An example of this is shown in Table 14.10.

Table 14.10 Word equations and balanced symbol (formula) equations.

<table>
<thead>
<tr>
<th>Type of equation</th>
<th>Word equation</th>
<th>Balanced equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Equation</td>
<td>nitrogen + hydrogen → ammonia</td>
<td>N₂ + 3H₂ → 2NH₃</td>
</tr>
<tr>
<td>What it tells us</td>
<td>Nitrogen reacts with hydrogen to form ammonia</td>
<td>One molecule of nitrogen (N₂) reacts with three molecules of hydrogen (H₂) to form two molecules of ammonia (NH₃)</td>
</tr>
</tbody>
</table>

In a balanced equation, the total number of atoms of each element on both sides of the equation must be the same. This is because atoms cannot be created or destroyed. In the equation for the reaction between nitrogen and hydrogen above, there are two nitrogen atoms and six hydrogen atoms in both the reactants and products.

You are often required to write a balanced equation. Here are some steps to follow plus two examples.

Step 1 Write the word equation.

Step 2 Rewrite the equation with formulae (be very careful to ensure the formulae are correct).

Step 3 Count the number of atoms of each element on each side of the equation. If they are the same then the equation is already balanced and nothing more needs to be done.

Step 4 If the equation is not balanced, then add in extra molecules to try and balance it. You must never change the formulae themselves. For example, you could not change the formula of water from H₂O to H₄O in example 1 to balance the H atoms because it is water that is formed and that has the formula H₂O and not H₄O.

Step 5 Write out the final balanced equation.
### Examples

**Table 14.11**

<table>
<thead>
<tr>
<th>Example 1</th>
<th>Example 2</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Step 1</strong></td>
<td>methane + oxygen → carbon dioxide + water</td>
</tr>
<tr>
<td><strong>Step 2</strong></td>
<td>CH₄ + O₂ → CO₂ + H₂O</td>
</tr>
<tr>
<td><strong>Step 3</strong></td>
<td>reactants products</td>
</tr>
<tr>
<td>C = 1</td>
<td>Al = 1</td>
</tr>
<tr>
<td>H = 4</td>
<td>O = 6</td>
</tr>
<tr>
<td>O = 2</td>
<td>N = 1</td>
</tr>
<tr>
<td>The equation is not balanced</td>
<td>The equation is not balanced</td>
</tr>
<tr>
<td><strong>Step 4</strong></td>
<td>Add another H₂O to the products (so there are now 2H₂O) to balance the H atoms:</td>
</tr>
<tr>
<td>CH₄ + O₂ → CO₂ + 2H₂O</td>
<td>Add two more HNO₃ to the reactants (so there are now 3HNO₃) to balance the N atoms</td>
</tr>
<tr>
<td>reactants products</td>
<td>reactants products</td>
</tr>
<tr>
<td>C = 1</td>
<td>Al = 1</td>
</tr>
<tr>
<td>H = 4</td>
<td>O = 6</td>
</tr>
<tr>
<td>O = 2</td>
<td>N = 1</td>
</tr>
<tr>
<td>Then add another O₂ to the reactants (so there are now 2O₂) to balance the O atoms:</td>
<td>Then add two more H₂O to the products (so there are now 3H₂O) to balance the O and H atoms:</td>
</tr>
<tr>
<td>CH₄ + 2O₂ → CO₂ + 2H₂O</td>
<td>Al(OH)₃ + 3HNO₃ → Al(NO₃)₃ + 3H₂O</td>
</tr>
<tr>
<td>reactants products</td>
<td>reactants products</td>
</tr>
<tr>
<td>C = 1</td>
<td>Al = 1</td>
</tr>
<tr>
<td>H = 4</td>
<td>O = 6</td>
</tr>
<tr>
<td>O = 4</td>
<td>N = 3</td>
</tr>
<tr>
<td>The equation is now balanced</td>
<td>The equation is now balanced</td>
</tr>
<tr>
<td><strong>Step 5</strong></td>
<td>CH₄ + 2O₂ → CO₂ + 2H₂O</td>
</tr>
</tbody>
</table>

Balanced equations sometimes include state symbols to show the state of each substance:

- (s) = solid
- (l) = liquid
- (g) = gas
- (aq) = aqueous (dissolved in water).

For example, the equation:

\[
\text{CaCO}_3(s) + 2\text{HCl(aq)} \rightarrow \text{CaCl}_2(aq) + \text{H}_2\text{O(l)} + \text{CO}_2(g)
\]

means that calcium carbonate solid reacts with an aqueous solution of hydrochloric acid to form an aqueous solution of calcium chloride, water liquid and carbon dioxide gas.
10 Magnesium reacts with sulfuric acid as shown:
\[
\text{Mg(s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{MgSO}_4(\text{aq}) + \text{H}_2(\text{g})
\]

a) What does the (s) mean?  

b) What does the (aq) mean?  

c) What does the (g) mean?  

11 Balance the following equations.

a) \( \text{K} + \text{I}_2 \rightarrow \text{KI} \)  
b) \( \text{Na} + \text{H}_2\text{O} \rightarrow \text{NaOH} + \text{H}_2 \)  
c) \( \text{CuCO}_3 \rightarrow \text{CuO} + \text{CO}_2 \)  
d) \( \text{Al} + \text{O}_2 \rightarrow \text{Al}_2\text{O}_3 \)  
e) \( \text{Ca} + \text{HCl} \rightarrow \text{CaCl}_2 + \text{H}_2 \)  
f) \( \text{KOH} + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + \text{H}_2\text{O} \)  
g) \( \text{C}_3\text{H}_2 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \)  
h) \( \text{H}_3\text{PO}_4 + \text{NaOH} \rightarrow \text{Na}_3\text{PO}_4 + \text{H}_2\text{O} \)  
i) \( \text{NH}_3 + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4 \)  
j) \( \text{NO} + \text{H}_2\text{O} + \text{O}_2 \rightarrow \text{HNO}_3 \)

12 Write a balanced equation for each of the following reactions.

a) sodium + oxygen \( \rightarrow \) sodium oxide  
b) propane \( (\text{C}_3\text{H}_8) \) + oxygen \( \rightarrow \) carbon dioxide + water  
c) calcium + water \( \rightarrow \) calcium hydroxide + hydrogen  
d) chlorine + sodium bromide \( \rightarrow \) sodium chloride + bromine  
e) magnesium oxide + nitric acid \( \rightarrow \) magnesium nitrate + water

Ionic equations

In a solid ionic compound, the positive and negative ions are bonded together strongly in a lattice. When it dissolves in water, the ions separate and become surrounded by water molecules (Figure 14.1).

When ionic compounds dissolved in water react, it is usual for some of the ions not to react and remain unchanged in the water. These are often called spectator ions as they are present but do not take part in the reaction.

We can write ionic equations for reactions involving ions. These ionic equations only show what happens to the ions that react. We do not include the spectator ions in these ionic equations.

The overall electric charge of the ions in the reactants must equal the overall electric charge of the ions in the products. This can sometimes be useful to help you check that the ionic equation is balanced or actually to help you balance the ionic equation.

Reaction of acids with alkalis

When sulfuric acid reacts with sodium hydroxide solution, it is only the hydrogen ions from the sulfuric acid and the hydroxide ions from the sodium hydroxide that react. These hydrogen ions and hydroxide ions...
Ionic equations

react to form water (Figure 14.2). The sulfate ions and sodium ions remain unchanged as they do not react and are left out of the ionic equation as they are spectator ions.

![Figure 14.2 Sulfuric acid reacting with sodium hydroxide.](image)

The ionic equation for this reaction can be written as:

$$\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(l)$$

from the acid from the alkali

When any acid reacts with any alkali, the ionic equation is the same. Some examples are shown in Table 14.12.

<table>
<thead>
<tr>
<th>Examples</th>
<th>What reacts</th>
<th>Ions that do not react</th>
<th>Ionic equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>sulfuric acid (aq) + sodium hydroxide (aq)</td>
<td>H(^+) ions from H(_2)SO(_4) OH(^-) ions from NaOH</td>
<td>SO(_4^{2-}) ions from H(_2)SO(_4) Na(^+) ions from NaOH</td>
<td>H(^+(aq)) + OH(^-(aq)) → H(_2)O(l)</td>
</tr>
<tr>
<td>hydrochloric acid (aq) + potassium hydroxide (aq)</td>
<td>H(^+) ions from HCl OH(^-) ions from KOH</td>
<td>Cl(^-) ions from HCl K(^+) ions from KOH</td>
<td>H(^+(aq)) + OH(^-(aq)) → H(_2)O(l)</td>
</tr>
<tr>
<td>nitric acid (aq) + calcium hydroxide (aq)</td>
<td>H(^+) ions from HNO(_3) OH(^-) ions from Ca(OH)(_2)</td>
<td>NO(_3^-) ions from HNO(_3) Ca(^{2+}) ions from Ca(OH)(_2)</td>
<td>H(^+(aq)) + OH(^-(aq)) → H(_2)O(l)</td>
</tr>
</tbody>
</table>
Displacement reactions

Ionic equations can also be written for displacement reactions that take place when a more reactive metal displaces a less reactive metal from a metal compound.

For example, when zinc reacts with copper sulfate solution, the zinc atoms in the zinc metal react with the copper ions in the copper sulfate (Figure 14.3). The sulfate ions are spectator ions and so do not appear in the ionic equation.

\[ \text{Zn(s)} + \text{Cu}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Cu(s)} \]

This and some other examples of displacement reactions are shown in Table 14.13. In these reactions it helps to use the electric charges of the ions to balance the equation. For example, in the second example in the table, two Ag\(^+\) ions are needed giving an overall 2+ charge on the left side of the equation to balance with the 2+ charge on the Fe\(^{2+}\) ion on the right side of the equation.

Table 14.13 Some displacement reactions.

<table>
<thead>
<tr>
<th>Examples</th>
<th>What reacts</th>
<th>Ions that do not react</th>
<th>Ionic equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>zinc (s) + copper sulfate (aq)</td>
<td>Zn atoms in Zn metal Cu(^{2+}) ions from CuSO(_4)</td>
<td>SO(_4^{2-}) ions from CuSO(_4)</td>
<td>Zn(s) + Cu(^{2+})(aq) → Zn(^{2+})(aq) + Cu(s)</td>
</tr>
<tr>
<td>iron (s) + silver nitrate (aq)</td>
<td>Fe atoms in Fe metal Ag(^+) ions from AgNO(_3)</td>
<td>NO(_3^-) ions from AgNO(_3)</td>
<td>Fe(s) + 2Ag(^+)(aq) → Fe(^{2+})(aq) + 2Ag(s)</td>
</tr>
<tr>
<td>aluminium(s) + copper chloride (aq)</td>
<td>Al atoms in Al metal Cu(^{2+}) ions from CuCl(_2)</td>
<td>Cl(^-) ions from CuCl(_2)</td>
<td>2Al(s) + 3Cu(^{2+})(aq) → 2Al(^{3+})(aq) + 3Cu(s)</td>
</tr>
</tbody>
</table>
Half equations

Many chemical reactions involve the transfer of electrons and half equations can be written for these reactions. These equations show the number of electrons that are gained or lost.

In these half equations:

- positive ions gain electrons
  (e.g. a 3+ ion gains 3 electrons, e.g. $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$)
- negative ions lose electrons
  (e.g. a 2– ion loses 2 electrons, e.g. $\text{S}^{2–} – 2\text{e}^- \rightarrow \text{S}$).

Half equations for the loss of electrons can be written in two ways. The electrons can be shown being taken away from the left-hand side, or shown on the right-hand side with the products.

For example, the half equation for the loss of electrons from $\text{S}^{2–}$ ions to form S can be written as:

$$\text{S}^{2–} – 2\text{e}^- \rightarrow \text{S} \quad \text{or} \quad \text{S}^{2–} \rightarrow \text{S} + 2\text{e}^-$$

Some elements contain diatomic molecules (e.g. $\text{H}_2$, $\text{O}_2$, $\text{Cl}_2$, $\text{Br}_2$, $\text{I}_2$). When balancing half equations that produce these elements, two ions are needed to make one diatomic molecule.

For example, when $\text{H}_2$ is formed from $\text{H}^+$ ions, two $\text{H}^+$ ions are needed which both gain one electron and so two electrons are gained altogether:

$$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$$

Test yourself

13 Write an ionic equation (including state symbols) for each of the following reactions between acids and alkalis.
   a) hydrochloric acid (aq) + sodium hydroxide (aq)
   b) nitric acid (aq) + potassium hydroxide (aq)
   c) sulfuric acid (aq) + calcium hydroxide (aq)
   d) phosphoric acid (aq) + sodium hydroxide (aq)

14 Write an ionic equation (including state symbols) for each of the following displacement reactions.
   a) Displacement of copper from copper(II) sulfate (aq) by magnesium.
   b) Displacement of silver from silver nitrate (aq) by magnesium.
   c) Displacement of zinc from zinc sulfate (aq) by aluminium.

15 Write an ionic equation (including state symbols) for each of the following reactions.
   a) Reaction of barium hydroxide (aq) with hydrochloric acid (aq).
   b) Displacement of silver from silver nitrate (aq) by zinc (s).
   c) Reaction of sulfuric acid (aq) with potassium hydroxide (aq).
   d) Displacement of nickel from nickel(II) sulfate (aq) by zinc (s).
When \(O_2\) is formed from \(O^{2-}\) ions, two \(O^{2-}\) ions are needed which both lose two electrons and so four electrons are lost altogether:

\[
2O^{2-} - 4e^- \rightarrow O_2 \quad \text{or} \quad 2O^{2-} \rightarrow O_2 + 4e^- 
\]

In half equations, the total electric charge on the left-hand side must equal the total electric charge on the right-hand side of the equation. This can be used to check that the half equation is balanced.

In this example, both the left and right-hand sides of the equation add up to the same overall charge (which is 0 in this case):

\[
\begin{align*}
\text{total charge:} & \quad 3^+ - 3^- = 0 \\
\text{left-hand side} & \quad \text{right-hand side}
\end{align*}
\]

Oxidation takes place when a substance loses electrons. Reduction takes place when a substance gains electrons. See the section on oxidation and reduction in Chapter 12. One way to remember whether electrons are lost or gained in oxidation and reduction is the phrase OIL RIG (see Figure 12.7).

**Electrolysis**

Half equations can be written for the reactions that take place at each electrode in electrolysis. Some examples are shown in Table 14.14.

<table>
<thead>
<tr>
<th>Substance (in molten state)</th>
<th>Negative electrode reaction</th>
<th>Positive electrode reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Copper sulfide</td>
<td>(Cu^{2+}) ions gain electrons to form (Cu)</td>
<td>(Cu^{2+} + 2e^- \rightarrow Cu)</td>
</tr>
<tr>
<td>Sodium chloride</td>
<td>(Na^+) ions gain electrons to form (Na)</td>
<td>(Na^+ + e^- \rightarrow Na)</td>
</tr>
<tr>
<td>Aluminium oxide</td>
<td>(Al^{3+}) ions gain electrons to form (Al)</td>
<td>(Al^{3+} + 3e^- \rightarrow Al)</td>
</tr>
</tbody>
</table>

**Displacement reactions**

Two half equations can be written for displacement reactions.

For example, in the reaction where zinc displaces copper from copper sulfate solution the overall ionic equation is:

\[
Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu
\]
The two half equations for this are:

Zn atoms lose electrons to form Zn\(^{2+}\) ions:
\[\text{Zn} - 2e^- \rightarrow \text{Zn}^{2+} \quad \text{(or Zn} \rightarrow \text{Zn}^{2+} + 2e^-)\]

Cu\(^{2+}\) ions in CuSO\(_4\) gain electrons to form Cu atoms:
\[\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}\]

For example, in the reaction when copper displaces silver from silver nitrate solution the overall ionic equation is:
\[\text{Cu} + 2\text{Ag}^+ \rightarrow \text{Cu}^{2+} + 2\text{Ag}\]

The two half equations for this are:

Cu atoms lose electrons to form Cu\(^{2+}\) ions:
\[\text{Cu} - 2e^- \rightarrow \text{Cu}^{2+} \quad \text{(or Cu} \rightarrow \text{Cu}^{2+} + 2e^-)\]

Ag\(^+\) ions in AgNO\(_3\) gain electrons to form Ag atoms:
\[\text{Ag}^+ + e^- \rightarrow \text{Ag}\]

Test yourself

16 Write a balanced half equation for each of the following conversions.
   a) \(\text{Mg}^{2+} \rightarrow \text{Mg}\)
   b) \(\text{Se}^{2-} \rightarrow \text{Se}\)
   c) \(\text{K}^+ \rightarrow \text{K}\)
   d) \(\text{Br}^- \rightarrow \text{Br}_2\)
   e) \(\text{O}^{2-} \rightarrow \text{O}_2\)
   f) \(\text{H}^+ \rightarrow \text{H}_2\)

17 Write two half equations to show what happens in the following displacement reactions.
   a) displacement of copper from copper(II) sulfate (aq) by magnesium
   b) displacement of silver from silver nitrate (aq) by magnesium
   c) displacement of zinc from zinc sulfate (aq) by aluminium
We are concerned that energy reserves are running out. What does the future hold? How will we generate electricity for the projected world population of 10 billion people in 2050?

This chapter covers specification points 6.1.1.1 to 6.1.3 and is called Energy. It covers energy changes in a system, the ways energy is stored before and after such changes, conservation and dissipation of energy and national and global energy resources.
Energy changes in a system, and the ways energy is stored before and after such changes

Energy stores and systems
We can begin to understand energy by studying changes in the way energy is stored when a system changes. A ‘system’ is an object or a group of objects that interact. Here are some situations with which you should be familiar.

● Throwing an object upwards
When you throw a ball upwards, just after the ball leaves your hand it has a store of kinetic energy. When the ball reaches its highest point, it has a store of gravitational potential energy. Just before you catch it again, it has a store of kinetic energy.

● Boiling water in a kettle
When you turn on your electric kettle, the water in the kettle gets hotter. There is now more internal (or thermal) energy stored in the hot water than there was in the cold water.

● Burning coal
When we burn coal there is a chemical reaction. Coal has a store of chemical energy which is transferred to thermal energy as it burns. A coal fire can warm up a room.
A car using its brakes to slow down
A moving car has a store of kinetic energy. When the car slows to a halt, it has lost this store of kinetic energy. The brakes exert a frictional force on the wheels, and the brakes get hot. The store of kinetic energy in the car has been transferred to a store of thermal energy in the brakes. This energy is then transferred to the surroundings.

Dropping an object which does not bounce
Just before the object hits the ground, it has a store of kinetic energy. After the object has stopped moving, the kinetic energy has been transferred to a store of internal energy in the object and the surroundings. So the object and the surroundings warm up a little. (You might hear a noise, but the energy carried by the sound is also transferred to the internal energy of the surroundings.)

Accelerating a ball with a constant force
We have a store of chemical potential energy in our muscles. When we throw a ball, our store of chemical potential energy decreases, and the ball’s store of kinetic energy increases. The hand applies a force to the ball and does work to accelerate it.

Energy stores
In the simple everyday events and processes that were described above, we identified objects that had gained or lost energy. For example, objects slow down or get hotter. We saw that the way energy is stored changes.

We use the following labels to describe the stores of energy you will meet:
- kinetic
- chemical
- internal (or thermal)
- gravitational potential
- magnetic
- elastic potential
- nuclear.

Counting the energy
Energy is a quantity that is measured in joules, J. Large quantities of energy are measured in kilojoules, kJ, and megajoules, MJ.

\[ 1 \text{ kJ} = 1000 \text{ J} (10^3 \text{ J}) \]
\[ 1 \text{ MJ} = 1000000 \text{ J} (10^6 \text{ J}) \]

The reason that energy is so important to us is that there is always the same energy at the end of a process as there was at the beginning. If we add up the total energy in all the stores, that number stays the same.

Figures 15.2 and 15.3 show some examples of counting the energy.

The principle of conservation of energy
The principle of conservation of energy states that the amount of energy always remains the same. There are various stores of energy. In any process energy can be transferred from one store to another, but energy cannot be destroyed or created.
Energy changes in a system, and the ways energy is stored before and after such changes

Transferring energy from one store to another

Light, sound and electricity are useful, but they are not stores of energy. They are ways of transferring energy from one store to a different energy store. You cannot go into a shop to buy a box of ‘electrical energy’, but you can buy a cell or battery. In a circuit, the chemical energy stored in a cell or battery causes electric charge to flow.

In a torch, the chemical energy stored in the battery causes an electric current (a flow of charge). The electric current causes the temperature of the bulb to increase so much that the bulb lights up. The light cannot be stored but it is useful. When the light strikes an object and is absorbed, the internal energy of the object increases.

If we drop a bunch of keys onto a table, the collision will make the air vibrate and we hear a sound. The sound wave transfers energy; it is not an energy store. The energy will transfer to the air and surrounding objects causing an increase in their store of internal energy.

Test yourself

1. Describe the energy stored in each of the following:
   a) a moving bicycle
   b) a compressed spring
   c) a bowl of breakfast cereal
   d) a rock lifted off the ground.

2. Explain what is wrong with this statement:
   ‘A car battery stores electrical energy for the lights, horn and starter motor.’

3. Describe how the stores of energy change from the beginning to the end of the following processes.
   a) A catapult launches a marble.
   b) A ball rolls along the ground and comes to rest.
   c) A butane gas camping cooker heats up a pan of water.
   d) A lump of soft putty falls to the ground.

4. Figure 15.4 shows a ball falling. Copy the diagram and fill in the values for the ball’s kinetic energy and gravitational potential energy at each height.

<table>
<thead>
<tr>
<th>Height (m)</th>
<th>Potential Energy (J)</th>
<th>Kinetic Energy (J)</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>90</td>
<td>0</td>
</tr>
<tr>
<td>B</td>
<td></td>
<td>30</td>
</tr>
<tr>
<td>C</td>
<td>30</td>
<td></td>
</tr>
<tr>
<td>D</td>
<td>0</td>
<td></td>
</tr>
</tbody>
</table>

Figure 15.4

Show you can...

Complete this task to show you understand the different stores of energy.

Name four stores of energy, and describe three examples of how energy can be transferred from one store to another.
Calculating the energy

In this section you will learn how to calculate the amount of energy associated with a moving object, a stretched spring and an object raised above the ground. These calculations are useful to us. For example, we can show how the energy in a system is redistributed when a change happens to the system.

Kinetic energy

The kinetic energy stored by a moving object can be calculated using the equation:

\[
\text{kinetic energy} = \frac{1}{2} \times \text{mass} \times (\text{speed})^2
\]

Where energy is in joules, J
mass is in kilograms, kg
speed is in metres per second, m/s.

Elastic potential energy

The amount of elastic potential energy stored in a stretched spring can be calculated using the equation:

\[
\text{elastic potential energy} = \frac{1}{2} \times \text{spring constant} \times (\text{extension})^2
\]

Where energy is in joules, J
spring constant is in newtons per metre, N/m
extension is in metres, m.

The spring constant, \( k \), is a measure of the spring’s stiffness; \( k \) is equal to the force needed to stretch the spring one metre.

The extension of the spring is the increase in its length from its original unstretched length.

Gravitational potential energy

The amount of gravitational potential energy gained by an object raised above ground level can be calculated using the equation:

\[
\text{gravitational potential energy} = m \times g \times h
\]

Where energy is in joules, J
mass is in kilograms, kg
gravitational field strength is in newtons per kilogram, N/kg
height is in metres, m.
Example 1
A crate has a mass of 80 kg. A crane lifts the crate from a height of 3 m above ground to a height 18 m above the ground. Calculate the increase in gravitational potential energy of the crate \( g = 10 \text{ N/kg} \).

\[ F = 784 \text{ N} \]
\[ W = 784 \text{ N} \]

\[ \Delta \text{Figure 15.5 A crane does work to increase the gravitational potential energy of the crate.} \]

**Answer**

Increase in height = 15 m

\[ \text{Increase in } E_p = mgh \]
\[ = 80 \times 9.8 \times 15 \]
\[ = 11 760 \text{ J or 12 kJ} \text{ (to 2 significant figures)} \]

Example 2
A car has a kinetic energy store of 64 800 J. It is travelling at a speed of 12 m/s. Calculate the mass of the car.

\[ E_k = \frac{1}{2} mv^2 \]
\[ 64 800 = \frac{1}{2} \times m \times (12)^2 \]
\[ m = \frac{2 \times 64800}{(12)^2} \]
\[ m = 900 \text{ kg} \]

Example 3
A spring has a force constant of 60 N/m. The spring is extended by 5 cm. Calculate the elastic potential energy stored in the spring.

\[ E_e = \frac{1}{2} kx^2 \]
\[ = \frac{1}{2} \times 60 \times (0.05)^2 \]
\[ = 0.075 \text{ J} \]

Remember to change the extension of 5 cm into metres.

Example 4
Calculate the change in kinetic energy stored when a car of mass 1200 kg slows down from 30 m/s to 20 m/s.

\[ \text{Change in kinetic energy} = \frac{1}{2} mv_1^2 - \frac{1}{2} mv_2^2 \]
\[ = \frac{1}{2} \times 1200 \times 30^2 \]
\[ - \frac{1}{2} \times 1200 \times 20^2 \]
\[ = 540000 - 240000 \]
\[ = 300000 \text{ J or 300 kJ} \]

Test yourself

5. Calculate the kinetic energy of a bullet of mass 0.015 kg, travelling at a speed of 240 m/s.
6. Calculate the increase in the gravitational potential energy store of a boy of mass 50 kg after he has climbed the Taipei 101 Tower, which is 440 m high.
7. A car has a mass of 1500 kg. It accelerates from a speed of 15 m/s to a speed of 20 m/s. Calculate the increase in kinetic energy of the car.
8. A car suspension spring has a spring constant of 2000 N/m. Calculate the elastic potential energy stored in the spring when it is compressed by 8 cm.
9. A meteor has a mass of 0.05 kg and it is travelling towards Earth at a speed of 30 km/s. Calculate the kinetic energy of the meteor.
Changes in energy

In this section you will learn how to use the energy equations to make predictions about changes to a system: which energy is transferred from one type of store to another.

Example 1
A ball of mass 100 g is thrown vertically upwards with a speed of 15 m/s. What is the maximum height the ball reaches?

Answer
As the ball rises, work is done by gravity to slow the ball down. Energy is transferred from the kinetic energy store of the ball to the gravitational potential energy store of the ball.

The principle of the conservation of energy tells us that the kinetic energy stored when the ball is at the bottom of its path is equal to the potential energy stored when the ball is at the top of its path, (if we assume that none of the energy is transferred to the surroundings).

\[ \frac{1}{2} mv^2 = mgh \]
\[ \frac{1}{2} \times 0.1 \times 15^2 = 0.1 \times 9.8 \times h \]
So \[ h = \frac{11.25}{0.98} = 11.5 \text{ m} \]

Example 2
A stretched bow stores 64 J of elastic potential energy. The bow fires an arrow of mass 20 g. Calculate the speed of the arrow as it leaves the bow.

Answer
As the bow does work to speed up the arrow, energy is transferred from the bow’s store of elastic potential energy to the arrow’s store of kinetic energy.

\[ \frac{1}{2} mv^2 = 64 \]
\[ \frac{1}{2} \times 0.02 \times v^2 = 64 \]
\[ v^2 = 6400 \]
\[ v = 80 \text{ m/s} \]

Example 3
A car of mass 1200 kg is parked on a 1 in 5 slope. The handbrake is released and the car rolls down the slope. When the car has travelled 20 m down the slope, how fast is it travelling?

Answer
A 1 in 5 slope means that the car goes up (or down) 1 m in height, when it travels 5 m along the slope. So, when the car has travelled 20 m along the slope it has gone down by a height of 4 m. The gravitational potential energy transferred can be calculated from this height.

\[ mgh = \frac{1}{2} mv^2 \]
Gravitational field strength, \( g = 9.8 \text{ N/kg} \)
\[ 1200 \times 9.8 \times 4 = \frac{1}{2} \times 1200 \times v^2 \]
\[ \frac{1}{2} v^2 = 39.2 \]
\[ v^2 = 78.4 \]
\[ v = 8.9 \text{ m/s (to 2 significant figures)} \]
Energy changes in a system, and the ways energy is stored before and after such changes

Test yourself
You can do these questions for practice, but you can also set up some experiments like this in the laboratory.

10 A spring with a spring constant of 200 N/m is fixed, so that it points vertically on to a table by using some Blu-Tack. The spring is compressed by 1 cm.
   a) Calculate the elastic potential energy stored in the spring.
      A polystyrene ball of mass 0.5 g is held on the spring and launched vertically upwards.
   b) i) How much gravitational potential energy does the ball have when it reaches its highest point?
      ii) Calculate the height the ball reaches when it is released.
      Assume the ball travels vertically upwards.

11 In Figure 15.8 a trolley is attached to a spring as shown. The trolley is pulled back to extend the spring by 15 cm. The spring has a spring constant of 80 N/cm.
   a) Calculate the elastic potential energy stored in the spring.
   b) i) State the kinetic energy stored in the trolley just after the spring reaches its unstretched length.
      ii) Calculate the maximum speed of the trolley.

12 In Figure 15.9 a trolley is attached to a mass of 0.2 kg. The 0.2 kg mass is allowed to fall to the floor.
   a) Calculate the gravitational potential energy of the 0.2 kg when it is 0.9 m above the ground.
   b) i) State the kinetic energy of the 0.2 kg mass and the trolley together, just before the mass hits the ground.
      ii) Calculate the maximum speed of the trolley, just as the mass hits the floor.

Show you can...
State the principle of the conservation of energy. Describe two demonstrations you have seen that help to explain this principle.

KEY TERM
Work When a force causes an object to move, work = force × distance

Work
In this section you are introduced to the definition of work, because by doing work we can transfer energy from one store to another.
A force does work on an object when the force causes the object to move, in the direction of the force. Work can be calculated using the equation:

\[ W = F \times s \]

*work* = *force* × *distance moved in the direction of the force*

Where work is in joules, J
force is in newtons, N
distance is in metres, m.

One joule of work is done when a force of 1 newton causes a displacement of 1 metre.

1 joule = 1 newton – metre

When we do work, by applying a force to move an object, we change the energy store of that object.

- When 200 J of work is done to lift a box upwards, the gravitational potential energy store of the box increases by 200 J.
- When 3000 J of work is done to accelerate a car, the kinetic energy store of the car increases by 3000 J.
- When 2 J of work is done to stretch a spring, the spring stores 2 J of elastic potential energy.

Electrical work is done by a battery when the battery makes a charge flow. (Current and charge is covered further in Chapter 16.)

**Example**

We can use the idea of work to help us calculate the braking distance of a car.

A car of mass 1500 kg is travelling at a speed of 20 m/s. The brakes apply a force of 5000 N to slow down and stop the car.

Calculate the braking distance of the car.

**Answer**

The decrease in the kinetic energy store of the car (transferred to the internal energy store in the brakes) = work done by the brakes.

\[ \frac{1}{2}m v^2 = Fs \]

\[ \frac{1}{2} \times 1500 \times 20^2 = 5000 \times s \]

\[ s = \frac{300000}{5000} \]

\[ s = 60 \text{ m} \]

**Power**

Often when we want to do a job of work, we want to do it quickly. We say that a crane that lifts a crate more quickly than another crane lifting the same crate is more powerful.
**Energy changes in a system, and the ways energy is stored before and after such changes**

**Power** is defined as the rate at which energy is transferred or the rate at which work is done. Power can be calculated by using these equations.

\[
P = \frac{E}{t} \quad \text{or} \quad P = \frac{W}{t}
\]

where power is in watts, \( W \)
energy transferred is in joules, \( J \)
time is in seconds, \( s \)
work done is in joules, \( J \)

An energy transfer of 1 joule per second is equal to a power of 1 watt.

Large powers are also measured in kilowatts, kW, and megawatts, MW.

\[
1 \text{ kW} = 1000 \text{ W (10}^3 \text{ W)} \quad 1 \text{ MW} = 1 000 000 \text{ W (10}^6 \text{ W)}
\]

**Example**

A weight lifter lifts a mass of 140 kg a height of 1.2 m in 0.6 s. Calculate the power developed by the weight lifter.

**Answer**

The potential energy transferred = \( mgh \)
\[E_p = 140 \times 9.8 \times 1.2 = 1646.4 \text{ J} \]

\[\text{power} = \frac{\text{energy transferred}}{\text{time}} = \frac{1646.4}{0.6} = 2700 \text{ W or 2.7 kW (to 2 significant figures)}\]

**Measuring your own power**

Work out your personal power by running up a flight of steps.

You need to know your mass, the time it takes you to run up the stairs and the vertical height of the stairs. Remember \( g = 9.8 \text{ N/kg} \).

1. Record your time and calculate the increase in your gravitational potential energy store.
2. Now calculate your power.
3. Explain why you need the vertical height of the staircase and not the length along the staircase.
Test yourself

13 What is the unit of power?

14 What is the connection between the energy transferred and power?

15 A crane lifts a weight of 12000 N through a height of 30 m in 90 s. Calculate the power output of the crane in kW.

16 Two students have an argument about who is more powerful. Peter says he is more powerful because he is bigger. Hannah says she is more powerful because she is quicker.

To settle the argument, they run up stairs of height 4.5 m. Use the information about their weights and times to settle the argument.

Table 15.1

<table>
<thead>
<tr>
<th></th>
<th>Weight in N</th>
<th>Fastest time in s</th>
</tr>
</thead>
<tbody>
<tr>
<td>Peter</td>
<td>760</td>
<td>3.80</td>
</tr>
<tr>
<td>Hannah</td>
<td>608</td>
<td>3.04</td>
</tr>
</tbody>
</table>

17 When an express train travels at a speed of 80 m/s, the resistive forces acting against it add up to 150 kN.

a) Calculate the work done against the resistive forces in 1 s.

b) Calculate the power output of the train, travelling at 80 m/s.

Show you can...

Design an experiment to measure the power of your arm as you lift a weight.

Energy changes in systems

You will find that some of this section – specifically, specific heat capacity – is also covered in Chapter 17.

The amount of energy stored or released from a system, as its temperature changes, can be calculated using the equation:

\[ \Delta E = mc \Delta \theta \]

change in thermal energy = mass \times specific heat capacity \times temperature change

Where change in thermal energy is in joules, J
mass is in kilograms, kg
specific heat capacity is in joules per kilogram per degree Celsius, J/kg °C
temperature change is in degrees Celsius, °C.

The specific heat capacity of a substance is the amount of energy required to raise the temperature of one kilogram of the substance by one degree Celsius.

The specific heat capacity varies from substance to substance.
An investigation to measure the specific heat capacity of a material

There are several different ways to obtain the data needed to calculate the specific heat capacity of a material. All of the methods involve the same idea; the decrease in one energy store (or work done) leads to an increase in the temperature of the material.

In this method energy is transferred from an electrical immersion heater to a metal block. The increase in the temperature of the metal block depends on the mass of the block and the specific heat capacity of the block.

Method
1. Measure the mass of the metal block \( m \) in kilograms.
2. Put the thermometer and immersion heater into the holes in the block.
3. Connect the immersion heater, joulemeter and power supply together as shown in Figure 15.12.
4. Measure the temperature of the metal block \( \theta_1 \) and then switch on the power supply.
5. Wait until the temperature of the block has gone up by about 10 °C then switch off the power supply. Write down the reading on the joulemeter \( E \). This gives you the amount of energy transferred to the immersion heater.
6. Do not take the immersion heater out of the block. Keep looking at the temperature and write down the highest temperature shown by the thermometer \( \theta_2 \).

If a joulemeter is not available set up the circuit shown in Figure 15.13. Use the following method to measure the energy transfer.

Switch the power supply on and again wait for the temperature of the block to increase by about 10 °C.

Watch the voltmeter and ammeter and write down the readings \( V \) and \( I \). The readings may change a little as the block gets warmer. Switch the power supply off and write down how many seconds the power supply was on for \( t \).

The energy transferred to the block can be calculated using the equation:

\[ E = V I t \]

Analysing the results
1. Calculate the increase in temperature of the block \( \Delta \theta = (\theta_2 - \theta_1) \).
2. Use the following equation to calculate the specific heat capacity \( c \) of the metal block:

\[ c = \frac{E}{m \Delta \theta} \]

3. It is likely that the value you calculate for the specific heat capacity of the metal will not be accurate.

**KEY TERM**

**Accurate** A measurement or calculated value that is close to the true value.
Look up the true value. How close is the true value to your experimental value? Calculate the difference between the two values. Do you think your experimental value is accurate?

4 If other people in the class have used different types of metal, then compare the different specific heat capacity values with the temperature rise of the metal. If you do this, it is important that the blocks used by everyone have the same mass and that the same amount of energy is transferred to each immersion heater.

You should find that the higher the temperature rise, the smaller the specific heat capacity of the metal.

**Taking it further**
1 Repeat the investigation, but this time cover the side of the block in a thick layer of insulating material.
2 Calculate a second value for specific heat capacity. Is this second value any more accurate? If it is, suggest why.

**Questions**
1 Why is it better not to remove the immersion heater from the block as soon as the heater is switched off?
2 The values calculated for specific heat capacity from this investigation would usually be greater than the true value. Explain why.

**Demonstration experiment**

Your teacher might demonstrate how you can work out the temperature of a hot object.

In Figure 15.14a) a small piece of steel is heated until it is red hot in a Bunsen flame. The steel is then quickly transferred to an insulated beaker, which contains 0.1 kg (100 ml) of water as shown in Figure 15.14b). ([Caution, the steel is red hot, and will burn the bench if dropped. The water will ‘spit’ as the steel is put into the beaker. Do not use a piece of steel of mass more than about 20 g.]

**Table 15.2 Example data and calculation**

<table>
<thead>
<tr>
<th>Specific heat capacity of water</th>
<th>4200 J/kg °C</th>
</tr>
</thead>
<tbody>
<tr>
<td>Specific heat capacity of steel</td>
<td>450 J/kg °C</td>
</tr>
<tr>
<td>Mass of steel</td>
<td>20 g</td>
</tr>
<tr>
<td>Temperature of water at the start of the experiment</td>
<td>19 °C</td>
</tr>
<tr>
<td>Temperature of water after the hot steel has been placed in it</td>
<td>34 °C</td>
</tr>
</tbody>
</table>

The thermal energy transferred by the hot steel equals the thermal energy gained by the water.

The thermal energy gained by the water

\[ \text{Thermal energy} = (mc\Delta t)_{\text{water}} \]

\[ = 0.1 \times 4200 \times 15 \]

\[ = 6300 \text{ J} \]
Conservation and dissipation of energy

You have already learnt that energy is conserved: energy cannot be created or destroyed. However, when energy is transferred from a source, it is not all transferred usefully. Often when energy is transferred some of the energy is dissipated or ‘wasted’.

When petrol is used to power a car, much of the energy is wasted. Only about a quarter of the chemical energy stored in the petrol is used to drive the car forwards (by doing work) and about three quarters of the chemical energy is wasted by heating up the engine.

Reducing energy dissipation

Every day in our lives we use energy from fuels for transport or for heating our homes. In all cases the chemical energy stored in the fuels is eventually transferred to the thermal energy store of the surrounding...
area. This is a less useful way of storing energy; the energy is being ‘wasted’. However, we try to ensure that as much energy is transferred usefully as possible. We try to minimise the amount of wasted energy.

Power stations
The purpose of a power station is to generate electricity, which can do useful work to light and heat our homes. Engineers design the generators in the power station to reduce the amount of waste energy in the power station. Generators are large machines which can dissipate energy by heating or by unwanted mechanical vibrations.

Car design
When we drive a car we want to make sure that as much of the chemical energy stored in the fuel as possible does useful work for us.

- Engineers design fuel efficient cars, which dissipate less energy.
- The car is made streamlined to reduce air resistance on the car.
- Moving parts of the car are lubricated with oil to reduce friction.

Keeping warm at home
When we heat our homes, the energy stored inside the house is dissipated through the roof, walls, windows or doors of the house to warm up the air outside. We want to make sure the energy escapes as slowly as possible.

Here are some ways in which we reduce unwanted energy dissipation at home.

Chimneys
Figure 15.15 shows a coal fire burning in a sitting room. Some of the energy from the burning coal is transferred to the air outside the house. This is wasted energy. By having the chimney inside the house, thermal energy can be transferred into the bedrooms upstairs. This is useful energy.

Walls
The rate at which energy is transferred through the walls of a house depends on four factors.

- The temperature difference between inside and outside. (Our heating bills are larger in winter than in summer.)
- The area of the walls. (Large houses cost more to heat than small houses.)
- The thermal conductivity of the walls. Some materials conduct heat well, metals, for example. These materials have a high thermal conductivity. Brick and glass are not good thermal conductors. They have relatively low thermal conductivities, but energy still flows out of a warm house through the walls and windows. The higher the thermal conductivity of a material, the higher the rate of energy transfer by conduction across the material.
- The thickness of the walls (or windows) is important. The thicker the walls, the slower the rate of energy loss.
Modern houses are built with two layers of brick as shown in Figure 15.17. Then the house is insulated with cavity wall insulation, between the two layers of brick. The foam which insulates the walls is full of trapped air. The air is a good insulator; it has a much lower thermal conductivity than brick or glass.

Loft insulation and carpets
The most efficient way to reduce energy loss from our house is to insulate the loft. A thick layer of loft insulation reduces energy loss through the roof (Figure 15.18). We also use insulating carpets to reduce energy loss through the floor.

The tiles on the kitchen floor feel cold when you walk on them in bare feet. These tiles are much better thermal conductors than carpets.

Double glazing
A thin pane of glass in a window transfers energy out of the house. We use double glazing to reduce energy loss through the windows. A layer of gas trapped between two panes of glass provides good insulation (Figure 15.19).

Efficiency
Efficiency is a way of expressing the proportion of energy that is usefully transferred in a process as a number. The most efficient machines transfer the highest proportions of input energy to useful output energy.
To calculate efficiency we use the equation:

\[
\text{efficiency} = \frac{\text{useful output energy transfer}}{\text{total input energy transfer}}
\]

Efficiency may also be calculated using the equation:

\[
\text{efficiency} = \frac{\text{useful power output}}{\text{total power input}}
\]

Efficiency is a ratio of energies or powers. So efficiency has no unit. We write efficiencies as a decimal or a percentage.

Example 1
A steam engine uses coal as its source of energy. When the chemical energy store of the coal in the engine’s furnace goes down by 150 kJ, the engine does 18 kJ of useful work against resistive forces. Calculate the efficiency of the engine.

Answer

\[
\text{efficiency} = \frac{\text{useful output energy transfer}}{\text{total input energy transfer}} = \frac{18}{150} = 0.12 \text{ or } 12\%
\]

Note: here both energies were expressed in kJ. Make sure you use the same unit on the top and bottom of the fraction.

Example 2
A joulemeter records that 18.2 J of electrical work is done when an electric motor lifts a 0.3 kg load through a distance of 0.90 m. Calculate the efficiency of the motor.

Answer

\[
\text{efficiency} = \frac{\text{useful output energy transfer}}{\text{total input energy transfer}} = \frac{\text{gain in } E_p}{\text{total input energy transfer}} = \frac{mgh}{W} = \frac{0.3 \times 9.8 \times 0.90}{18.2} = \frac{18.2}{18.2} = 0.15 \text{ or } 15\% \text{ (2 significant figures)}
\]

Increasing the efficiency of an intended energy transfer
Whenever we do a job of work, we want to ensure that as much energy as possible is transferred usefully, and that little energy is dissipated wastefully.

When we move an object we transfer energy by doing work, and work is calculated using the equation:

\[
W = Fs
\]
Conservation and dissipation of energy

where \( F \) is the force applied and \( s \) is the distance moved in the direction of the force. We do less work if the forces of friction or air resistance that act against us are small. When frictional forces act, energy is transferred to the thermal energy store of the surroundings. This is wasteful.

We reduce friction by:
- using wheels
- applying lubrication.

We reduce air resistance by:
- travelling slowly
- streamlining.

When we streamline a car (for example), we are shaping its surface so that air flows past the car and offers as little resistance to the motion of the car as possible.

Increasing efficiency using machines

We can also increase the efficiency of a job by using a machine. Figure 15.21 shows two men lifting a load of bricks on a building site. The man on the ground pulls with a force of 250 N. Because there are four ropes in their system of pulleys, the ropes apply a force of \( 4 \times 250 \text{ N} \) (1000 N), which is enough to lift the bricks and the pulley, and to overcome friction.

Without the machine the men would have to carry the bricks up the ladder in smaller loads. This means that they would have to work to carry the bricks and to lift their own weight. The machine allows them to apply a smaller force to lift the bricks, but they have to pull the rope 4 times as far. Using the machine is far more efficient than climbing up and down the ladder. It saves time and is much safer.

250 N
\[ 4 \times 250 \text{ N} = 1000 \text{ N} \]
5 m
2800 N

Test yourself

26 Define efficiency.
27 A student compares two machines, A and B. Both machines transfer the same input energy. The student discovers that machine A wastes less energy than machine B. Which machine is more efficient?
28 Phil is in the gym doing pull-ups (Figure 15.22). Each time he does a pull-up his store of chemical energy decreases by 1500 J. Phil’s mass is 72 kg and he lifts himself up 0.5 m in one pull-up.
   a) Calculate the gravitational potential energy stored after one pull-up.
   b) Calculate the efficiency of Phil’s body during this exercise.
29 A car’s engine is supplied with one kilogram of fuel, which stores 45 MJ of chemical energy. The efficiency of the car is 36%.
   Calculate the amount of energy available for useful work against resistive forces.
30 Why does a streamlined car use fuel more efficiently than another similar car which experiences larger air resistance forces?
31 a) Why does the machine in Figure 15.21 allow the men to work more efficiently?
   b) Suggest another machine that increases the efficiency of a job.
   How does the machine ensure that more energy is usefully transferred?
Every day we depend on various energy resources to make our lives comfortable. A hundred and fifty years ago our ancestors walked to school, lived in cold houses and did not have electricity in their homes. Now we travel by car, train or bus, we live in warm homes, and all of us use electricity to run the many appliances we have at home.

- **Fossil fuels**
  Much of our energy in the UK comes from the fossil fuels: coal, oil and gas. Like all fuels, they store energy. However, to release the energy, the fossil fuels must be burned. Once the fuels are burnt, they are gone forever, because fossil fuels have taken millions of years to be formed. Fossil fuels are described as non-renewable energy resources, because there is a finite supply of them (once gone, they cannot be replaced).

By contrast renewable energy resources will never run out. We can obtain renewable energy from the Sun, tides, waves and rivers, from the wind and from the thermal energy of the Earth itself.

- **Using fuels**
  In the UK the three main fossil fuels provide most of the energy for our needs. These needs include:
  - **Transport.** Fuels such as petrol, diesel and kerosene are produced from oil. These fuels drive our cars, trains and planes. Electricity is also used to run our trains, and cars are being developed to run from electricity supplies. In 50 years’ time, we may not be able to fly on holiday.
  - **Heating.** Most of our home heating is provided by gas and electricity. Gas pipes run into our houses to provide energy to our boilers. Some homes are warmed by oil-fired boilers or by burning solid fuels such as coal and wood.
Electricity. In the UK, electricity is generated using different energy resources, but most of our electricity is generated by burning fossil fuels. Table 15.3 shows the percentage of electricity generated using different energy resources. Gas and coal provide a convenient and relatively cheap way to generate our electricity. However, the burning of gas and coal produces carbon dioxide, which most scientists think is responsible for global warming. In November 2015, the UK government announced that coal fuelled power stations will be phased out by 2025. More electricity will be generated by gas, which produces less carbon dioxide than coal.

Table 15.3 Percentage of UK electricity generated by different energy resources.
(Source: Department of Energy and Climate Change, September 2015.) We can expect a greater amount of electricity to be generated by renewable sources in future.

<table>
<thead>
<tr>
<th>Energy resource used to generate electricity</th>
<th>Percentage of UK electricity generated from each energy resource 2015</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gas</td>
<td>30</td>
</tr>
<tr>
<td>Coal</td>
<td>20</td>
</tr>
<tr>
<td>Nuclear</td>
<td>22</td>
</tr>
<tr>
<td>Biomass</td>
<td>9</td>
</tr>
<tr>
<td>Wind</td>
<td>11</td>
</tr>
<tr>
<td>Hydroelectric</td>
<td>2</td>
</tr>
<tr>
<td>Oil</td>
<td>2</td>
</tr>
<tr>
<td>Solar</td>
<td>4</td>
</tr>
</tbody>
</table>

Fossil fuels and acid rain
One of the products of burning coal is sulfur dioxide. When sulfur dioxide combines with water, acid rain is produced. Acid rain damages buildings and kills plants.

Sulfur dioxide can be removed from the waste gases of burning coal, but this is expensive.

Power stations
The average consumption of electrical power in the UK is about 36 GW and our peak consumption (in the evening) reaches 57 GW. (1 GW is 1 gigawatt, which is 1000 million watts or 10^9 watts.)

Figure 15.25 shows the layout in a coal-fired power station. You are not expected to remember this for your GCSE but it is helpful to understand the principle behind generating electricity: an energy resource such as gas or water drives a turbine; the turbine then drives the generator, which makes electricity.
Due to concerns over global warming, which is caused (in part) by the burning of fossil fuels, governments must find alternative energy resources. Climate change meetings attended by governments usually involve agreements to reduce fossil fuel use.

Nuclear power

Nuclear power generates about 22% of the electricity in the UK. There are currently plans for a new nuclear power station at Moorside in West Cumbria. The Moorside power station will begin generating electricity in 2024. The peak generating capacity of the power station will be 3.4 GW.

The nuclear fuels used are mainly uranium and plutonium. These are also non-renewable energy resources. However, nuclear fuel contains a huge amount of nuclear energy, and it is estimated that there is enough uranium to last thousands of years.

Nuclear power has the advantage of producing no pollutant gases. However, we need to be very careful how we store nuclear waste.

Biomass

Waste products can provide fuel for some small electrical generators. Much of the waste is wood, so this is a renewable energy resource. Many of the biomass generators are also used to heat factories or houses directly.

Biofuels emit carbon dioxide when they are burnt. However, the plants that become biofuels used carbon dioxide as they grew. So, overall biofuels do not add to the amount of carbon dioxide in the atmosphere. They are 'carbon neutral'.

Using tides

Every day tides rise and fall. Massive amounts of water move in and out of river estuaries. It is estimated that the energy of the tides could generate about 20% of Britain’s electricity.

At present there is little electricity generated by tidal energy. Expensive barriers must be constructed. The largest tidal barriers in the world generate about 250 MW of electricity.

A barrage is like a dam built across a river estuary (Figure 15.26). The barrage has underwater gates that open as the tide comes in and then close to keep the water behind the barrage (Figure 15.27). When the tide goes out, a second set of gates is opened. Water flows out of these gates and drives turbines that are connected to generators as it does so.
Hydroelectric power

Figure 15.28 Hydroelectric power stations can be huge and generate vast amounts of power.

Hydroelectric power stations generate about 2% of Britain's electricity, but they generate about 10% of the world's electricity.

Many hydroelectric power schemes have had environmental impacts, as a new lake is formed.

- Forests have been cut down.
- Farmland is lost.
- Wildlife habitats have been destroyed.
- Many people have had to move homes.

Wind power

When the wind blows with sufficient force, the blades of a wind turbine rotate. The blades turn a generator which produces electricity. Wind power has the advantage of being environmentally clean and is a renewable energy source. There are no waste products.

However, wind power also has some disadvantages.

- Wind power is unreliable. If the wind is too light, little power is produced. If the wind is too strong, generators can overheat, so the blades have to be stopped from moving.
- Some people think that wind farms are unsightly and spoil the look of the countryside.
- Wind generators make a low frequency sound, which creates noise pollution. Some people find it most unpleasant living close to a wind farm.

At present 11% of Britain's electricity comes from wind power. The government's target is to increase this figure to about 20%. Although wind power is unreliable, when the wind does blow, we can turn off gas generators. So wind power will save fossil fuels and reduce greenhouse gas emissions.
Solar power
Solar cells use energy directly from the Sun to generate electricity. Solar cells generate electricity on a small scale.

In some countries solar electricity generation is **unreliable**, due to the weather, but it can still make a useful contribution to a country’s overall power supply and does not contribute to **global warming** and is **renewable**.

Geothermal energy
A small number of countries are able to make good use of geothermal energy to generate electricity. Iceland generates 30% of its electricity by taking advantage of the volcanic activity on the island. In some cases hot water is used directly to warm houses. Hot water or steam is also used to generate electricity. This form of electricity generation has the advantage of being **environmentally clean** and renewable.

Test yourself

32 a) What is a non-renewable energy resource? Give one example.
   b) What is a renewable energy resource? Give one example.

33 Name a common fuel used in a nuclear power station. Is this fuel renewable or non-renewable?

34 In Britain 20% of our electricity is generated in coal-fired power stations.
   a) What are the advantages of using coal to generate electricity?
   b) State two environmental problems caused by using coal to generate electricity.

35 a) What environmental problems are caused by building a hydroelectric power station?
   b) Give one advantage of hydroelectric power.

36 Why are tides a more reliable way of generating electricity than wind power?

37 Britain plans to have 12,000 wind turbines spread from the south to the north of the country. Why does spreading out the wind farms increase the reliability of wind power?
38 A wind turbine is designed to produce a maximum power of 4 MW. However, due to variations of wind speed, the generator only produces this power for 10% of the time. How many such wind turbines are required to replace a coal-fired power station which generates 2000 MW of power all the time?

39 Figure 15.33 shows the layout of a pumped storage power station. Water from the high level lake generates electricity by flowing through the turbines which are connected to generators. These are placed above the low level lake. When there is a low demand for electricity, the generators are driven in reverse to pump water back into the high level lake. This ensures there is enough water to generate electricity next time the demand is high.

a) Why is this sort of power station useful to electricity companies?

b) Does this power station produce any pollution or greenhouse gases
   i) when generating electricity, ii) when pumping water back up the hill?

c) Use the information in Figure 15.33 to calculate the gravitational potential energy transferred per second when the generators are working.

d) The generators are 80% efficient. Calculate the power output of the power station in MW.

Figure 15.34 shows the typical power use in the UK on a spring day.

e) At what time of the day does the pumped storage station i) generate electricity, ii) pump water back up the hill? Give reasons for your answer.

Show you understand the issues that affect the generation of electricity by writing a plan for electricity generation in the UK for 2040. Which types of generation are you going to use, and where will you site your power stations? Your brief report should give proper consideration to ethical issues.
Energy: a summary

In this book we provide clear definitions using ‘key terms’ where possible. No such easy definition exists for energy. Instead we have included this brief summary to pull the ideas about energy together.

Energy is an idea that cannot be described by a single process, nor is energy something we can hold or measure directly. However, we pay an enormous amount of attention to energy because it is conserved. There are many different stores of energy. In any process, energy can be transferred from one store to another store, but energy is never destroyed or created.

Energy stores

Stores of energy include:

- kinetic
- chemical
- internal (or thermal)
- gravitational potential
- magnetic
- elastic potential
- nuclear.

Energy transfers

There are various ways that energy can be transferred:

- by mechanical work
- by electrical work
- by heating
- by radiation.

The word ‘radiation’ includes light and all electromagnetic waves. Radiation also includes ‘mechanical radiation’ such as sound and shock waves. Sometimes people refer to light, sound and electrical energy. However, light, sound and electricity are not energy stores, they transfer energy from one store to another. Here are some examples.

A hot cup of tea has a store of thermal energy. As the tea cools, it heats the surroundings. Energy is transferred to the thermal store of the surroundings and the temperature of the surroundings goes up (but is too small to measure).

A battery stores chemical energy. Energy is transferred by electrical work to the lamp. The lamp does not store energy. The lamp transfers energy to the thermal store of the surroundings by heating (conduction and convection) and radiation. Some of this radiation is useful to us as it is transferred; this is visible light.
A motor is used to lift a mass. A battery stores chemical energy. Energy is transferred by electrical work from a battery to a motor. The motor does not store energy. The motor does mechanical work and transfers energy to the gravitational potential store of the raised mass. The motor also transfers energy to the thermal store of the surroundings by heating and by making a noise.

**Figure 15.37** Energy transfer to and from an electric meter.

**And finally...**

Energy can be transferred from one energy store to a different store, and there is as much energy stored at the end of the process as there was at the beginning. The reason energy is useful to us is that it allows us to do various jobs of work and to keep warm.
Chapter review questions

1. Describe the energy store of each of the following:
   a) an electrical battery
   b) a moving car
   c) a stretched rubber band
   d) a lake of water behind a hydroelectric dam.

2. In the following processes energy is transferred from one energy store to another store or stores. State the stores of energy at the beginning and end of each process.
   a) A battery lights a lamp.
   b) A bowl of hot soup cools.
   c) A battery is connected to a motor which is lifting a load.
   d) A firework rocket has been launched into the air and is travelling upwards.

3. Your feet feel warmer on a carpet than they do a tiled kitchen floor. Give one reason why.

4. Calculate the energy stored in each of these examples.
   a) A car of mass 1400 kg travels at a speed of 25 m/s.
   b) A suspension spring for a truck with a spring constant of 40000 N/m is compressed by 5 cm.
   c) A suitcase of mass 18 kg is placed in a luggage rack 2.5 m above the floor of a train.

5. Figure 15.38 shows a girl on a slide. Her mass is 45 kg.
   a) Calculate the girl’s speed at the bottom of the slide. Ignore the effects of friction.
   b) Explain why the girl’s speed is likely to be less than the answer calculated in part (a).

6. A gymnast of mass 55 kg lands from a height of 5 m onto a trampoline (see Figure 15.39). Calculate how far the trampoline stretches before the gymnast comes to rest. The trampoline has a spring constant of 35000 N/m.

7. Some lead shot with a mass of 50 g is placed into a cardboard tube as shown in Figure 15.40. The ends of the tube are sealed with rubber bungs to keep the lead shot in place. The tube is rotated so that the lead shot falls and hits the bung at the bottom.
   a) Why does the temperature of the lead shot increase after it has fallen and hit the lower bung?
   b) A student rotates the tube 50 times. Calculate the total decrease in the gravitational potential energy store of the lead shot in this process.
   c) The specific heat capacity of lead is 160 J/kg °C. Calculate the temperature rise of the lead shot after the student has rotated the tube 50 times.
   d) In practice, the temperature rise of the lead shot is likely to be less than your answer in part (c). Give a reason why.

8. A girl kicks a football with a force of 300 N. The girl’s foot is in contact with the ball for a distance of 0.2 m. The ball has a mass of 450 g. Calculate the speed of the ball just after it has been kicked.
9 An electricity supply provides electrical power to a motor at a rate of 800 W. The motor lifts a crate of mass 80 kg through a height of 3 m in 12 seconds. Calculate the efficiency of the motor.

10 A student wants to calculate the specific heat capacity of a liquid. First he determines that the liquid has a density of 900 kg/m³.
   a) The student places a volume of 200 cm³ of liquid into an insulated beaker. Calculate the mass of the liquid in kg. [1 cm³ = 10⁻⁶ m³]
      The student measures the temperature of the liquid. It is 22 °C. The student then heats the liquid with a heater that has a power rating of 24 W. He heats the liquid for 10 minutes.
   b) Calculate the energy transferred to the liquid in 10 minutes.
   c) After 10 minutes the student finds the liquid has risen to a temperature of 72 °C. Calculate the specific heat capacity of the liquid.

11 A scientist observes a grasshopper as it jumps. The grasshopper takes 25 milliseconds to take off, and reaches a speed of 3.0 m/s. The grasshopper has a mass of 1.5 g.
   a) Calculate the kinetic energy of the grasshopper after its jump.
   b) Calculate the mechanical power developed by the grasshopper’s legs.
Practice questions

1 Which of the following is the correct unit for power?
   newtons joules watts [1 mark]

2 Which of the following is the correct unit for specific heat capacity?
   J/kg °C J kg/°C J kg °C [1 mark]

3 The energy input to a machine is 2000 J. The machine transfers 600 J of useful energy. Calculate the efficiency of the machine. [2 marks]

4 The British government has planned to build up to 12 000 wind generators. Give one advantage of wind power and one disadvantage. [2 marks]

5 Many of the world’s electricity power stations burn fossil fuels.
   a) Burning fossil fuels produces carbon dioxide.
      What effect can an increase in carbon dioxide levels have on the Earth’s atmosphere? [1 mark]
   b) Figure 15.41 shows how much carbon dioxide is produced for each unit of electricity generated in coal-, gas- and oil-burning power stations.
      i) Which type of fossil fuel produces the most carbon dioxide for each unit of electricity generated? [1 mark]
      ii) Why is a bar chart drawn to show the data and not a line graph? [1 mark]
   c) Biofuels are renewable energy resources that can be used to generate electricity.
      i) Name one other type of renewable energy resource. [1 mark]
      ii) Most biofuels are derived from plants. When biofuels are burned carbon dioxide is produced. Why does burning a biofuel have less effect on the atmosphere than burning coal? [1 mark]

6 An advertisement for solid fuel firelighters claims:
   Be certain of a fast fire ...
   ... use H&S, the firelighters that give out more energy than any others

   a) To test this claim a student plans the following investigation.
      i) What type of variable is the brand of firelighter? [1 mark]
      ii) Name two control variables in this investigation. [2 marks]
      iii) Give one experimental hazard in this investigation. [1 mark]
      iv) Suggest one change that the student could have made to improve the resolution of the temperature readings. [1 mark]
b) To compare the data the student drew the bar chart shown in Figure 15.44.

![Figure 15.44](image)

**Figure 15.44**

i) Was the data collected by the student sufficient to confirm the claim made by the maker of H&S firelighters? Give a reason for your answer. [2 marks]

ii) Give two reasons why the decrease in the chemical energy store of the firelighter is greater than the increase in the thermal energy store of the water. [2 marks]

7 The roof of the Tokyo Skytree is 495 m high and can be climbed using its steps. A tourist decided to climb the tower. He took 35 minutes to do it and he had a mass of 60 kg.

a) Calculate the increase in the gravitational potential energy store of the tourist in climbing the steps. [3 marks]

b) Calculate his average power output during the climb. [3 marks]

c) The energy to make the climb is transferred from the chemical energy store of the tourist. Eating one slice of bread provides 400 kJ of chemical energy. Calculate the number of slices of bread he should eat for breakfast to provide the energy for the climb.

(Assume his body is 20% efficient at transferring energy from his chemical energy store to gravitational potential energy.) [3 marks]

8 In Figure 15.45 a conveyor belt is used to lift bags of cement on a building site.

![Figure 15.45](image)

**Figure 15.45**

a) A 40 kg bag of cement is lifted from the ground to the top of the building. Calculate its gain in gravitational potential energy. [3 marks]

b) The machine lifts five bags per minute to the top of the building. Calculate the useful energy delivered by the machine each second. [2 marks]

c) The machine is 35% efficient. Calculate the input power to the machine while it is lifting the bags. [3 marks]

9 A car and its passengers have a combined mass of 1500 kg. The car is travelling at a speed of 15 m/s. It then increases speed to 25 m/s. Calculate the increase in kinetic energy of the car. [3 marks]

10 Figure 15.46 shows a pirate boat theme park ride which swings from A to B to C and back.

a) As the boat swings from A to B a child increases her kinetic energy store by 10830 J. The child has a mass of 60 kg and sits in the centre of the boat. Calculate the speed of the child as the boat passes through B. [3 marks]

![Figure 15.46](image)

**Practice questions**
b) Sketch a graph to show how the child’s gravitational potential energy changes as the boat swings from A to B to C. [3 marks]

c) Calculate the change in height of the ride.
(Assume the decrease in the gravitational potential energy store as the child falls is transferred to the kinetic energy store of the child.) [3 marks]

11 An electric winch is used to pull a truck up a slope, as shown in Figure 15.47.

a) How much work is done in lifting the truck 15 m? [3 marks]

The winch uses a 6 kW electrical supply, and pulls the truck up the slope at a rate of 5 m/s.

b) How long does it take to pull the truck up the slope? [2 marks]

c) How much work is done by the winch? [2 marks]

d) Calculate the efficiency of the winch. [2 marks]

12 A barrage could be built across the estuary of the River Severn. This would make a lake with a surface area of about 200 km² (200 million m²).

Figure 15.48 shows that the sea level could change by 9 m between low and high tide; but the level in the lake would only change by 5 m.

a) Calculate the gravitational potential energy transferred when a mass of 1 kg falls from A to B. [3 marks]

b) Calculate the number of cubic metres of water that flow out of the lake between high and low tides. [2 marks]

c) Calculate the mass of water that flows out between high and low tides. A cubic metre of water has a mass of 1 kg. [2 marks]

d) Use your answers to parts (a) and (c) to calculate how much energy can be transferred from the tide. Assume that position A is the average position of a cubic metre of water between high and low tide. [3 marks]

e) The time between high and low tide is approximately 6 hours. Use this figure to calculate the average power available from the dam. Give your answer in megawatts. [3 marks]

f) What are the advantages and disadvantages of the Severn barrage as a possible source of power? [4 marks]
Working scientifically: Uncertainty, errors and precision

When an electric motor is used to lift a weight, the power supply does electrical work to make the motor turn.

The motor transfers some energy usefully. This increases the gravitational potential energy stored by the weight. The rest of the energy is dissipated to the surroundings.

Susan decided to find out if the efficiency of an electric motor depends on the size of the weight being lifted. To do this she set up the apparatus shown in Figure 15.49.

Susan started with a 2N weight. She switched on the power supply and increased the potential difference (p.d.) until the motor just lifted the weight from the floor to the bench top; a distance of 0.8 m. The joulemeter recorded the energy transfer to the motor. Susan repeated this step twice more. Her results are recorded in the table.

<table>
<thead>
<tr>
<th>Trial</th>
<th>Joulemeter reading</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>14.8</td>
</tr>
<tr>
<td>2</td>
<td>15.3</td>
</tr>
<tr>
<td>3</td>
<td>14.9</td>
</tr>
</tbody>
</table>

Show that the mean (average) joulemeter reading was 15.0.

The three readings taken from the joulemeter are all close to the mean. The values for the energy transferred are precise.

When the same quantity is measured several times, the bigger the spread of the measurements about the mean, the less precise the measurements are. The precision of a set of measurements depends on the extent of the random errors.

The uncertainty in these energy values is ±0.3. The uncertainty is worked out by calculating the difference between the mean value and the value furthest away from the mean (in this case 15.3 – 15.0).

KEY TERMS

**Precise** A set of measurements of the same quantity that closely agree with each other.

**Random error** An error unpredictable. Repeating the measurement and then working out a mean will reduce the effect of a random error. In a graph, data that is scattered about the line of best fit shows a random error.

**Uncertainty** For a set of measurements, the difference between the maximum value and the mean or the minimum value and the mean gives a measure of the uncertainty in the measurements.
An uncertainty in a set of data can be caused by random errors or a systematic error. In this investigation, judging when the weight reaches the bench top and then switching off the power supply is a random error. If whenever the joulemeter is reset it does not go back to zero, this is a systematic error. This type of systematic error is also called a zero error.

Susan repeated the procedure with a range of different weights. For each weight, she obtained three energy values and recorded the mean. These values are shown in the table.

<table>
<thead>
<tr>
<th>Weight lifted in N</th>
<th>Mean energy input to the motor in J</th>
<th>Percentage efficiency of the motor</th>
</tr>
</thead>
<tbody>
<tr>
<td>15.0</td>
<td>2</td>
<td>10.7</td>
</tr>
<tr>
<td>16.9</td>
<td>3</td>
<td>14.2</td>
</tr>
<tr>
<td>21.2</td>
<td>4</td>
<td></td>
</tr>
<tr>
<td>25.5</td>
<td>5</td>
<td>15.7</td>
</tr>
<tr>
<td>33.5</td>
<td>6</td>
<td>14.3</td>
</tr>
<tr>
<td>46.0</td>
<td>7</td>
<td>12.2</td>
</tr>
</tbody>
</table>

2 In Susan’s investigation:
   a) What range of weights was used?
   b) What was the independent variable?
   c) Which variable was controlled during the investigation?

3 How could Susan tell that the joulemeter she used did not have a zero error?

4 Copy and complete the table above by calculating the missing efficiency value.

5 Draw a graph of percentage efficiency against weight lifted.

6 What can you conclude from this investigation?
At the beginning of the 20th century, very few people had electricity supplied to their homes. Now, a century later, the supply of electricity to our homes, offices and streets is an essential part of life. However, electricity comes at a price and we must be careful how much we use it. We must plan ahead to make sure we are able to provide electricity far into the future.

This chapter covers specification points 6.2.1.1 to 6.2.4.3 and is called Electricity. It covers current, potential difference and resistance, series and parallel circuits, domestic uses and safety and energy transfers.
When two objects rub against each other, electrons can transfer from one object to another. When electrons transfer to an object it becomes negatively charged. When electrons leave an object it becomes positively charged.

Two like charges repel each other.

Two unlike charges attract each other.

An electrical current is a flow of charge.

In metals, current is a flow of electrons.

Some materials are good conductors of electricity.

Some materials do not conduct electricity. These are called insulators.

A cell or battery has a store of chemical energy. The energy stored decreases when a current flows.

When the same current passes through a number of components they are said to be in series.

In a parallel circuit, the current divides into different branches.

Test yourself on prior knowledge
1. Are the lights in your home in series or in parallel? How can you tell?
2. a) Name three good conductors of electricity.
   b) Name three good insulators of electricity.
A **cell** has a store of chemical energy. The longer line shows the positive terminal of the cell; the shorter line is the negative terminal.

A **battery** is two or more cells in series.

A **switch** breaks or rejoins the circuit, turning the current off or on. This switch is open so no current flows.

This **switch** is closed, so current can flow.

A **bulb**. When a bulb (or lamp) lights, it shows a current is flowing.

A **resistor**. The value of the resistor affects the size of the current.

An **ammeter** measures the current.

A **voltmeter** measures the size of the potential difference.

A **variable resistor** allows the current to be varied.

A **fuse** melts (blows) when the current is greater than the fuse’s current rating.

A **diode** only allows the current to flow in one direction—the direction of the arrows.

A **light-emitting diode (LED)** emits light when a current flows through it. An LED can be used as an indicator, to show when an appliance is switched on.

A **light-dependent resistor (LDR)**. The resistance of an LDR is low in bright light and higher when the light intensity is low.

A **thermistor**. The resistance of a thermistor changes with temperature. The resistance of a thermistor is low at high temperatures and high at low temperatures.

> Figure 16.3 Circuit symbols.

---

**Test yourself**

1. Which one of the following is the correct symbol for an LDR?

   - (a) ![Diagram](https://via.placeholder.com/150)
   - (b) ![Diagram](https://via.placeholder.com/150)
   - (c) ![Diagram](https://via.placeholder.com/150)

   ▲ Figure 16.4

2. Use words from the list below to label each of the components in the circuit in Figure 16.5.

   - cell
   - resistor
   - fuse
   - lamp
   - switch
   - diode

3. Draw a circuit diagram to show a cell in series with an ammeter, variable resistor and bulb.

   ▲ Figure 16.5

---

**Show you can...**

Complete this task to show that you understand how to draw and design electrical circuits. Draw a circuit diagram to show how two lamps can be connected to a battery, with components that allow the two lamps to be dimmed independently.
Current and charge

Figure 16.6 shows a circuit diagram. In this circuit a cell provides a potential difference of 1.5 V, giving a current of 0.1 A in the circuit.

The potential difference (or p.d.) is a measure of the electrical work done by a cell (or other power supply) as charge flows round the circuit. The potential difference is measured in volts (V). Here the cell provides a potential difference of 1.5 V. (Remember the positive terminal of the cell is shown with the long line and the negative terminal with a shorter line.)

It is quite common to call the potential difference ‘voltage.’ However, you will find that the examination papers (and this book) will use the term ‘potential difference’.

In a metal the current is carried by electrons which are free to move. The electrons are repelled from the negative terminal of the cell and attracted towards the positive terminal.

In a circuit the direction of the electric current is always shown as the direction in which positive charge would flow – from the positive terminal of the battery to the negative terminal. Current was defined in this way before the electron was discovered, at a time when people did not understand how a wire carried a current. So in Figure 16.6 the direction of current is shown from positive to negative.

The amount of charge flowing round in the circuit is measured in coulombs, C. One coulomb of charge is equivalent to the charge on 6 billion billion electrons.

The unit of current is the ampere, A. This unit is often abbreviated to amp. Small currents can be measured in milliamps (mA).

\[ 1 \text{ mA} = 0.001 \text{ A} \ (10^{-3} \text{ A}) \]

The current at all points of the circuit shown in Figure 16.6 is the same. So the two ammeters on either side of the lamp read the same current – in this case 0.1 A.

Current and the flow of charge are linked by the equation.

\[ Q = I \times t \]

charge flow = current \times time

where charge is in coulombs, C

current is in amps, A

time is in seconds, s.
In Figure 16.8c), the potential difference (voltage) of the cell is 1.5 V but now the resistance has been increased by adding a second resistor to the circuit. This makes the current smaller.

Controlling the current

You can change the size of the current in a circuit by changing the potential difference of the cell or battery, or by changing the components in the circuit.

In Figure 16.8a) a current of 1A flows. In Figure 16.8b) an extra cell has been added. Now the current is larger.

Test yourself

4 State the unit of each of the following quantities:
   a) potential difference
   b) current
   c) charge.

5 In Figure 16.8 the potential difference is 3.0V at point X. What is the current at points Y and Z?

6 a) A charge of 3C flows round a circuit in 2 seconds. Calculate the current.
   b) A torch battery delivers a current of 0.3A for 20 minutes. Calculate the charge which flows round the circuit.
   c) A thunder cloud discharges 5C of charge in 0.2 ms. Calculate the current.
   d) A mobile phone battery delivers a current of 0.1mA for 30 minutes. Calculate the charge which flows through the battery in this time.

Show you can...

Show you understand the nature of an electrical current, by explaining the relationship between a current and charge.

Example

In Figure 16.6 the current of 0.1 A flows for 30 minutes. How much charge flows round the circuit?

Answer

\[ Q = It \]

\[ Q = 0.1 \times 1800 \]

\[ = 180 \text{ C} \]

TIPS

- In a series circuit, increasing the potential difference increases the current.
- In a series circuit, increasing the resistance makes the current smaller. Resistance is a measure of a component’s opposition to the current.

KEY TERM

Resistor A component that acts to limit the current in a circuit. When a resistor has a high resistance, the current is low.
In Figure 16.9 a variable resistor has replaced the fixed resistor in the circuit. You can control the size of the current by adjusting the resistor.

- **Ammeters and voltmeters**

Figure 16.10 shows you how to set up a circuit using an ammeter and a voltmeter.

- The **ammeter** is set up in **series** with the resistor. The same current flows through the ammeter and the resistor.
- The **voltmeter** is placed in **parallel** (on a separate branch) with the resistor. The voltmeter measures the potential difference across (between the ends of) the resistor.
- The voltmeter only allows a very small current to flow through it, so it does not affect the current flowing around the circuit.

- **Resistance**

The **current** through a component depends on two things:

- the **resistance** of the component
- the **potential difference** across the component.

The circuit in Figure 16.10 can be used to determine the value of the resistor in the circuit.

The current, potential difference or resistance can be calculated using the equation:

\[ V = IR \]

**potential difference = current \times resistance**

where potential difference is in volts, V
- current is in amps, A
- resistance is in ohms, \( \Omega \)

Resistances are also measured in kilohms (kΩ) and megohms (MΩ).

\[ 1 \text{kΩ} = 1000 \Omega \quad (10^3 \Omega) \quad 1 \text{MΩ} = 1000000 \Omega \quad (10^6 \Omega) \]

It is very important that you can use this equation in each of its three forms. The triangle in Figure 16.11 is a useful way to remember it.

**Example**

A potential difference of 12 V is applied across a resistor of 240 Ω. Calculate the current through the resistor.

**Answer**

If you are in doubt about rearranging the formula on the examination paper, try to remember the triangle. Put your finger over the symbol you want to work out, in this case \( I \), and we get:

\[ I = \frac{V}{R} \]

\[ = \frac{12}{240} = 0.05 \text{ A} \]
Test yourself

7 You use a variable resistor to act as a dimmer for a torch lamp. How should you change the resistance to make the lamp brighter?

8 Copy and complete this table.

Table 16.1

<table>
<thead>
<tr>
<th>Electrical device</th>
<th>Potential difference across device in V</th>
<th>Current through device in A</th>
<th>Resistance of device in Ω</th>
</tr>
</thead>
<tbody>
<tr>
<td>Resistor</td>
<td>1.5</td>
<td>0.05</td>
<td>20</td>
</tr>
<tr>
<td>Lamp</td>
<td>230</td>
<td>0.05</td>
<td>23</td>
</tr>
<tr>
<td>Heater</td>
<td>230</td>
<td>0.04</td>
<td>75</td>
</tr>
<tr>
<td>LED</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Electric car motor</td>
<td>72</td>
<td>48</td>
<td></td>
</tr>
</tbody>
</table>

If you have an ohmmeter you can measure the resistance of the wire directly; you do not need to calculate it.

Analysing the results

1 Plot a graph of resistance against the length of the wire.

2 Your graph should give a straight line going through the origin (0, 0). If it does, then you have shown that the resistance is directly proportional to the length of the wire.

Taking it further

1 Use the same circuit and a range of wires of different cross-sectional area to show that the resistance of a wire is inversely proportional to the area of the wire.

2 Set up the circuit shown in Figure 16.10. Use the circuit to investigate the resistance of combinations of resistors in series and in parallel. (See the section on series and parallel circuits later in this chapter.) First take measurements to calculate the value of a single fixed value resistor. Add a second resistor in series with the first resistor. Measure the p.d. across both resistors and the current through the resistors. Use these values to calculate the resistance of the two resistors in series. Repeat the experiment but with the two resistors connected in parallel. What do your results show happens to the resistance of a circuit when a) resistors are connected in series b) resistors are connected in parallel?

Questions

1 What was the dependent variable in this investigation?

2 What aspects of the investigation were important in trying to stop the wire from getting hot?

3 The width of the crocodile clips makes it difficult to measure the exact length of wire connected into the circuit. What type of error will this cause?
**Current–potential difference characteristic graphs**

**An ohmic conductor**
For some resistors, at constant temperature, the current through the resistor is proportional to the potential difference across it. A graph of current against potential difference gives a straight line. If the direction of the p.d. is reversed, the graph has the same shape. The resistance is the same when the current is reversed.

The resistor in this case is said to be **ohmic**.

**A filament lamp**
The current–potential difference graph for a filament lamp does not give a straight line; the line curves away from the current axis (y-axis).

The current is not proportional to the applied potential difference. The lamp is a **non-ohmic** resistor.

As the current increases, the resistance gets larger. The temperature of the filament increases when the current increases. So we can conclude that the resistance of the filament increases as the temperature increases. Reversing the p.d. makes no difference to the way the resistance of the lamp changes. The resistance always increases when the temperature of the filament increases.

**KEY TERMS**

**Ohmic** The current flowing through an ohmic conductor is proportional to the potential difference across it. If the p.d. doubles, the current doubles. The resistance stays the same.

**Non-ohmic** The current flowing through a non-ohmic resistor is not proportional to the potential difference across it. The resistance changes as the current flowing through it changes.

**A diode**
A **diode** is a component that allows current to go in only one direction. For a forward potential difference, current starts to flow when the potential difference reaches about 0.7 V. When the potential difference is ‘reversed’, there is no current at all. The diode has a very high resistance in the reverse direction.

A **light-emitting diode** (LED) is a special type of diode that lights up when a current flows through it. This is useful because it allows an LED to be used as an indicator to show us that a small current is flowing.

**Changing resistance**
Some resistors change their resistance as they react to their surroundings. The resistance of a **thermistor** decreases as the temperature increases. You can control its temperature by putting it into a beaker of warm or cold water, as in Figure 16.16.
By gently heating the water, the resistance of the thermistor can be found at different temperatures (Figure 16.17). You could use an ohmmeter to measure the resistance directly.

A thermistor can be used as the sensor in a temperature-operated circuit, such as a fire alarm, so when the temperature and current increase to a certain level, the alarm sounds. Some electronic thermometers use a thermistor to detect changes in temperature. The change in the resistance of a thermistor can be used to switch on (or off) other electrical circuits automatically.

The resistance of a light-dependent resistor (LDR) changes as the light intensity changes. In the dark the resistance is high but in bright light the resistance of an LDR is low. This is shown in Figure 16.18. A higher current flows through the resistor in bright light because the resistance is lower.
LDRs can be used as sensors in light-operated circuits, such as security lighting. The change in resistance of LDRs is used in digital cameras to control the total amount of light that enters the camera.

Investigating the $I$–$V$ characteristic of a circuit component

1. Set up the circuit shown in Figure 16.19. A suitable power supply to use is four 1.5V cells joined in series.
2. Connect a 6V filament lamp into the circuit where it says component.
3. Adjust the variable resistor to give a potential difference (p.d.) of 1V across the lamp.
4. Write the readings on the voltmeter (p.d.) and the ammeter (current) in a suitable table.
5. Adjust the variable resistor so that you can obtain a set of p.d. and current values. Write the new values in your table.
6. Reverse the connections to the power supply. The readings on the voltmeter and ammeter should now be negative. Obtain a new set of data with the p.d. increasing negatively.

You could leave the variable resistor out of the circuit and change the p.d. and current by simply connecting across one cell, then two cells, then three cells and lastly all four cells.

Analysing the results

1. Plot a graph of current against potential difference ($I$–$V$ characteristic graph). Draw the axes so that you can show all of the data, the positive values and the negative values.
2. You should notice that plotting the negative values for p.d. and current gives the same shape graph line as plotting the positive values.
3. If you used a variable resistor you would have been able to increase the p.d. using a smaller interval than if you simply connected across 1, 2, 3 then 4 cells. The advantage of using smaller intervals is that you can be more confident that the shape you draw for your graph line is correct.

Taking it further

Replace the filament lamp with a low value resistor and then a diode. Obtain a set of p.d./current data for each component. Plot an $I$–$V$ characteristic graph for each component. Remember to obtain the data needed to plot the negative part of the graph.

Questions

1. The $I$–$V$ graph for a resistor (at constant temperature) is a straight line. Reversing the power supply does not change the shape of the graph line. What does this tell you about the resistance of the resistor and the direction of the current through the resistor?
2. How is the $I$–$V$ graph for a diode different from the $I$–$V$ graph for a filament lamp?
3. Why does having a small interval between values allow you to be more confident that the shape you draw for your graph line is correct?
4. If the p.d. were increased from 0V to 6V in 13 equal intervals, what would be the interval between p.d. values?

KEY TERM

Interval The difference between one value in a set of data and the next.
Test yourself

9 Use the graph in Figure 16.17 to find the resistance of the thermistor at temperatures of:
   a) 0°C
   b) 40°C
   c) 90°C.
10 What is meant by an ohmic resistor?
11 Figure 16.20 shows the current–potential difference graph for a filament bulb.
   a) Calculate the resistance of the filament when the applied voltage is:
      i) 1 V
      ii) 3 V.
   b) What causes the resistance to change?
12 Figure 16.18 shows a current–potential difference graph for an LDR in dim and bright light. Calculate the resistance of the LDR in:
   a) bright light
   b) dim light.
13 a) Draw a circuit diagram to show a cell, a 1 kΩ resistor and an LED used to show that there is a current flowing through the resistor.
   b) Draw a circuit diagram to show how you would investigate the effect of light intensity on the resistance of an LDR.

Show you can...

Show that you understand about various types of resistors, by explaining how the following behave in a circuit:
   a) an LDR
   b) an LED
   c) a thermistor.

Series and parallel circuits

Series circuits

In Figure 16.21 two identical bulbs are connected in series with a 12 V battery. The p.d. of 12 V from the battery is shared equally so that each bulb has 6 V across it.

Having two bulbs in the circuit rather than one increases the resistance, so the current decreases. The two bulbs in series will not be as bright as a single bulb.

The potential differences do not always split equally. In Figure 16.22 bulb 1 has a larger resistance than bulb 2. There is a larger potential difference across bulb 1 than across bulb 2. As the current flows, bulb 1 transfers more energy to the surroundings than bulb 2. Therefore bulb 1 is brighter than bulb 2.
### Parallel circuits

In Figure 16.24 two identical bulbs have been connected in parallel with the 12 V battery. Now there is a 12 V potential difference across each bulb, and the same current flows through each bulb.

When two bulbs are joined in parallel, the total current in the circuit increases. So the combined resistance of the two bulbs is less than either bulb by itself. The total current in the circuit is the sum of the currents through the two bulbs.

#### Parallel circuit rules

The rules for parallel circuits are as follows:

- the **potential difference** across each component is the **same**
- the **total current** through the whole circuit is the **sum** of the currents through the separate components
- the **total resistance** of two resistors in parallel is less than the resistance of the smaller individual resistor.

### Cells and batteries

A battery consists of two or more electrical cells. When cells are joined in series, the total potential difference of the battery is worked out by adding the separate potential differences together. This only works if the cells are joined facing the same way, positive (+) to negative (−) (see Figure 16.25).
Test yourself

14 Figure 16.26 shows a circuit with a cell, two ammeters and a resistor. What reading does the ammeter on the right-hand side give?

15 a) In Figure 16.27, what is the resistance between
   i) AB
   ii) CD?

b) Which of the following correctly states the resistance between E and F?

- 30 Ω
- more than 20 Ω
- less than 10 Ω
- between 20 Ω and 10 Ω

16 a) What is the potential difference of the cell in Figure 16.28 a)?
b) What is the potential difference across R_2 in Figure 16.28 b)?

17 State the values of A_1, A_2, A_3 and A_4 in Figures 16.29 a) and b).

18 Work out the potential difference of each cell combination shown in Figure 16.30.
Circuit calculations

You can use the series circuit rules to solve circuit problems. Two worked examples are given here.

Example

Use the information in Figure 16.31 to work out the resistance of the bulb.

\[ \text{Answer} \]

The potential difference across the bulb is \( 9V - 3V = 6V \)

\[ \text{So } R = \frac{V}{I} = \frac{6V}{0.3A} = 20\Omega \]

Example

The light dependent resistor in Figure 16.32 is in bright light. Use the information in the diagram to work out its resistance.

\[ \text{Answer} \]

\[ a) \] The total resistance of the series circuit can be calculated using:

\[ R = \frac{V}{I} = \frac{12}{0.06} = 200\Omega \]

Therefore the LDR’s resistance is:

\[ 200\Omega - 150\Omega = 50\Omega \]

OR you could work out the potential difference across the 150\( \Omega \) resistance:

\[ V = IR \]

\[ = 0.06 \times 150 \]

\[ = 9V \]

The potential difference across the LDR is:

\[ 12V - 9V = 3V \]

Then \[ R = \frac{V}{I} = \frac{3}{0.06} = 50\Omega \]

\[ b) \] When it gets dark the resistance of the LDR increases. So the current decreases. Therefore, the potential difference across the 150\( \Omega \) resistor drops, and the LDR gets a larger share of the potential difference.

\[ \text{Answer} \]

b) Explain what happens to the voltmeter reading when the light intensity drops.

\[ \text{Example} \]

Use the information in Figure 16.31 to work out the resistance of the bulb.

\[ \text{Answer} \]

The potential difference across the bulb is \( 9V - 3V = 6V \)

\[ \text{So } R = \frac{V}{I} = \frac{6V}{0.3A} = 20\Omega \]
Domestic use and safety

In the home we rely on electricity for heating, lighting, cooking, washing and powering devices which we use for work or leisure.

Direct and alternating potential difference

A cell or battery provides potential difference. A direct potential difference remains always in the same direction, and causes a current to flow in the same direction. This is a direct current (d.c.).

If an alternating power supply is used in a circuit, the potential difference switches direction many times each second. This is an alternating potential difference, which causes the current to switch direction. So an alternating current (a.c.) is one that constantly changes direction, passing one way around a circuit and then the other.

Figure 16.37 shows graphs of how a.c. and d.c. power supplies change with time. The d.c. supply remains constant at 6 V; the a.c. supply changes from positive to negative.

Note that the peak a.c. potential difference is a little higher than 6 V. This is to make up for the time when the potential difference is close to zero. A 6 V a.c. supply and a 6 V d.c. supply will light the same bulb equally brightly.

Mains supply

The mains electricity (domestic electricity supply) is supplied by alternating current. In the UK it has a potential difference of about 230 V and a frequency of about 50 Hz.
A frequency of 50 Hz means that the cycle shown in Figure 16.37 repeats itself 50 times per second – or one cycle takes 1/50 of a second.

In Figure 16.38, a 230 V a.c. supply provides current to a cooker. In diagram (a) the current goes one way round the circuit, then in diagram (b) the current is reversed. In each case energy is transferred to the cooker.

### Test yourself

22 a) What is meant by the terms a.c. and d.c.?

b) Explain the difference between direct and alternating potential difference.

23 In the USA the mains electricity supply is 115 V 60 Hz. Explain the difference between the mains electricity supply in the USA and the mains electricity supply in the UK.

### Cables

We use many electrical appliances at home, which we connect to a wall socket using a three-core cable and plug. The wires inside the cables connect the appliance to the plug and have a cross-sectional area of 2.5 mm². These cables should carry no more than a 13 A current. Appliances such as showers and cookers need larger currents. These appliances are connected to the mains supply using thicker cables.

**Live, neutral and earth wires**

The insulation covering the three wires inside a cable is colour coded so we know which is which.

- **Live wire** – *brown*
- **Neutral wire** – *blue*
- **Earth wire** – *green* and *yellow* stripes

The *live wire* carries the alternating potential difference from the mains supply. The neutral wire completes the circuit. So the live and neutral wires carry the current to and from an electrical appliance. The earth wire is there to stop an appliance becoming live; it only carries a current if there is a fault.

However, the three wires have different potentials.

- The earth wire is at 0 V. It only carries a current if there is a fault.
- The potential difference between the live wire and earth (0 V) is about 230 V. A bare live wire is a hazard even if it is not delivering a current to an appliance. If a person touches a bare live wire, the current will pass through the person’s body, giving an electric shock. In effect we would act as an earth.
- The neutral wire is close to earth potential (0 V). So touching the neutral wire would not give us an electric shock.

**TIP**

A live wire may be dangerous even when a switch in the mains circuit is open.
Energy and charge

When charge passes through a resistor, the resistor gets hot. This is because electrons collide with the atoms in the resistor as they pass through it. The atoms increase their thermal energy store and vibrate faster, making the resistor hotter. The energy in the chemical store of the battery decreases as work is done to move the electrons round the circuit. The temperature of the resistor goes up and the resistor also heats up the surroundings.

The amount of energy transferred by electrical work depends on two factors:

- the potential difference, $V$, across the resistor
- how much charge, $Q$, flows through the resistor.

The energy transferred by electrical work can be calculated using this equation:

$$E = VQ$$

energy transferred = potential difference $\times$ charge

where energy is in joules, J
potential difference is in volts, V
charge is in coulombs, C

Power and energy

When an electricity company sends a bill to a household, the company is charging for the energy transferred. So knowing how much energy is transferred by an electrical appliance is important.

The energy transferred by an electrical appliance depends on two things:

- the power of the appliance
- the time the appliance is switched on.

So, the energy transferred by electrical work can also be calculated using this equation:

$$E = Pt$$

energy transferred = power $\times$ time

where energy is in joules, J
power is in watts, W
time is in seconds, s

By rearranging this equation, the power of an appliance can be calculated.
We can combine the two equations as follows.

\[
P = \frac{VQ}{I}
\]

but since

\[
I = \frac{Q}{t}
\]

\[
P = VI
\]

\[\text{power} = \text{potential difference} \times \text{current}\]

where power is in watts, W
potential difference is in volts, V
current is in amps, A.

Since potential difference, \(V = IR\), this equation can also be written as:

\[
P = I^2 R
\]

\[\text{power} = \text{current squared} \times \text{resistance}\]

where power is in watts, W
current is in amps, A
resistance is in ohms, \(\Omega\).

Most electrical appliances have an information plate, which tells us the power of the appliance and the potential difference of the supply. From this you can work out the current the appliance will draw.

**Example 1**

The information plate on a convection heater is marked as follows:

230V 50Hz 1800W

Calculate the current the heater draws from the mains supply.

**Answer**

\[
P = VI
\]

\[
1800 = 230 \times I
\]

\[
I = \frac{1800}{230} = 7.8 \text{A}
\]

**Example 2**

A current of 4.7 A passes through a 30 \(\Omega\) resistor. Calculate the power transferred to the resistor.

**Answer**

\[
P = I^2 R
\]

\[
= 4.7^2 \times 30
\]

\[
= 663 \text{W}
\]
Figure 16.41 Electricity is transmitted through the National Grid using overhead transmission cables.

Figure 16.42 shows how we can reduce the current in the transmission cables.

- The generator in the power station sends a current of 1000 A at a potential difference of 25 000 V into the National Grid.
- A device called a step-up transformer steps the potential difference up to 400 000 V, but reduces the current to 62.5 A.
- With a current of 62.5 A, less energy is transferred into heating up the transmission cables than with a current of 1000 A.
- Near our homes, step-down transformers step the potential difference down to a safe level of 230 V, so a larger current is available to use in the home.
- This makes the National Grid system an efficient way to transfer energy.

**KEY TERMS**

- **Step-up transformer** A transformer that increases potential difference and decreases current.
- **Step-down transformer** A transformer that decreases potential difference and increases current.

**TIP**

Transformers make the transfer of energy from power stations to customers more efficient due to reduced energy losses.
16 Electricity

Show you understand about the transmission of electricity by completing this task.

a) Why is electricity transmitted across the country at very high potential differences?
b) Why must we use only low potential differences in our homes?

Test yourself

26 In Figure 16.42 the resistance of the power line is 200Ω.
   a) Calculate the power transferred in heating up the line when
      i) a current of 100 A passes through it
      ii) a current of 1000 A passes through it.
   b) Explain why transformers are used to reduce the current passing through the transmission cables of the National Grid.

27 A man decides to use a large cell in his house to supply current to light a 36 W lamp in his garden shed 10 m away (Figure 16.43). The resistances of the two wires to the shed are each 4Ω.
   a) Calculate the power transferred in each wire.
   b) Calculate the fraction of the power delivered by the cell which is used to light the bulb.

Show you can...

Show you understand about the transmission of electricity by completing this task.

a) Why is electricity transmitted across the country at very high potential differences?
b) Why must we use only low potential differences in our homes?
Chapter review questions

1. Which current in Figure 16.44 is the largest, A, B or C? Which is the smallest?

2. Figure 16.45 shows a simplified picture of the inside of a small fan heater. The electrical wiring is not shown. Draw, using circuit symbols, a diagram to show how the heating element, fan and switches would be connected together so that:
   - when the mains is switched on, the fan comes on
   - both heating elements can be switched on independently
   - on hot days the heater can be used as a cooling fan.

3. Figure 16.46 shows a simplified circuit diagram for the front lights of a car. The metal body acts as a wire for the circuit.
   a) Which switch operates the headlights?
   b) Can the sidelights be switched on without the headlights? Give a reason for your answer.
   c) If one headlight breaks, will the other still work? Give a reason for your answer.
   d) What change would you have to make to the circuit if the car’s body was made of plastic?

4. A 10 Ω resistor is placed in series with a bulb, a switch and a 9 V battery.
   a) Draw the circuit diagram.
   b) When the switch is closed a current of 0.3 A flows. Calculate:
      i) the power supplied by the battery
      ii) the power transferred to the resistor
      iii) the power transferred by the bulb.

5. In Figure 16.47 each cell has a potential difference of 1.5 V.
   a) What potential difference does the battery produce?
   b) State the reading on the voltmeter.
   c) Calculate the current through the ammeter.
   d) Calculate the resistance of i) the resistor, ii) the bulb.
6 In Figure 16.48 a 24 Ω resistor is placed in series with component X. Calculate the resistance of X.

\[ V = 12 \text{ V} \]
\[ I = 0.2 \text{ A} \]

\[ R = \frac{V}{I} = \frac{12 \text{ V}}{0.2 \text{ A}} = 60 \Omega \]

\[ \text{Figure 16.48} \]

7 A set of decorative lights has 115 identical bulbs connected in series. Each bulb is designed to take a current of 0.05 A. The set of bulbs is connected directly into the 230 V mains electricity supply.

a) What is the potential difference across one of the bulbs?

b) Calculate the resistance of one of the bulbs.

c) Calculate the resistance of all of the bulbs in series.

d) Calculate the power of the set of bulbs.

8 Explain why the resistance of a filament bulb increases as the current flowing through the bulb increases.

9 When the switch is closed in Figure 16.49, the bulb lights dimly at first. However, the bulb gets brighter slowly. Explain why.

\[ \text{Figure 16.49} \]
Practice questions

1 Figure 16.50 shows a simple circuit.

\[ \text{Figure 16.50} \]

\[ \text{A} \quad \begin{array}{c} 6 \text{V} \\ \downarrow \\ \downarrow \\ \downarrow \end{array} \]

\[ \text{0.15A} \]

\[ \text{40\,\Omega} \]

\[ \text{0.10A} \]

\[ \text{\textbf{a)}} \quad \text{The four cells are identical.} \\
\text{What is the potential difference of one cell? \quad [1 mark]} \\
\text{\textbf{b)}} \quad \text{State the reading on the voltmeter. \quad [1 mark]} \\
\text{\textbf{c)}} \quad \text{The current through the 40\,\Omega resistor is 0.15\,A.} \\
\text{The current through the bulb is 0.10\,A.} \\
\text{What is the reading on the ammeter? \quad [1 mark]} \\
\text{\textbf{d)}} \quad \text{Use the correct answer from the box to copy and complete the sentence.} \\
\text{less than} \quad \text{equal to} \quad \text{greater than} \\
\text{The bulb has a resistance} \quad 40\,\Omega \\
\text{Give a reason for your answer. \quad [2 marks]} \]

2 Figure 16.51 shows a simple circuit. The circuit includes an LDR.

\[ \text{Figure 16.51} \]

\[ \text{A} \quad \begin{array}{c} 3 \text{V} \\ \downarrow \\ \downarrow \end{array} \]

\[ \text{\textbf{a)}} \quad \text{How does the resistance of an LDR change with changing light intensity? \quad [1 mark]} \\
\text{\textbf{b)}} \quad \text{Figure 16.52 shows how the reading on the ammeter changes with light intensity.} \\
\text{\textbf{i)}} \quad \text{What is the current in the circuit when the light intensity is equal to the value marked ‘X’? \quad [1 mark]} \\
\text{\textbf{ii)}} \quad \text{Calculate the resistance of the LDR when the light intensity is equal to the value ‘X’. Give the unit. \quad [2 marks]} \\
\text{\textbf{iii)}} \quad \text{Suggest a practical use for this circuit. \quad [1 mark]} \]

3 Mains electricity provides an alternating current (a.c.); a cell provides a direct current (d.c.).

\text{Describe the difference between a.c. and d.c. \quad [2 marks]} \]

4 An electric iron has been wired without an earth connection. After years of use the live wire becomes loose and touches the metal part of the iron.

\text{\textbf{a)}} \quad \text{A man touches the iron and receives an electric shock.} \\
\text{Sketch a diagram to show the path taken by the current. \quad [1 mark]} \\
\text{\textbf{b)}} \quad \text{The mains potential difference is 230\,V. The man’s resistance is 46\,k\Omega.} \\
\text{Calculate the current that passes through the man. \quad [3 marks]}
5 Figure 16.54 shows three resistors connected to a 12 V battery.
   a) Calculate the currents through the ammeters $A_1$ and $A_2$. [2 marks]
   b) Which resistor has the greater value, $R_1$ or $R_2$. Give a reason for your answer. [2 marks]
   c) Calculate the resistance $R_1$. [3 marks]

6 A filament lamp is connected to a 12 V battery. The current through the lamp is recorded by a data logger when the lamp is switched on. Figure 16.55 shows how the current changes just after the lamp is switched on.

   a) Describe how the current changes just after the bulb is switched on. [2 marks]
   b) Use the graph to determine:
      i) the maximum current
      ii) the current after 1 second. [2 marks]
   c) The resistance of the filament increases as it gets hotter. Use this information to explain the shape of the graph. [3 marks]
   d) Use the graph to calculate the power of the bulb when it is working at its steady temperature. [3 marks]

7 Figure 16.56 shows a circuit diagram which includes a diode. Figure 16.57 shows how the current through the diode varies with the potential difference across the diode.

   a) A student sets up the circuit and measures the current through the diode as 20 mA. Use Figure 16.56 to determine the potential difference across the diode for this current. [1 mark]
   b) Calculate the potential difference across the 260 Ω resistor when the current through the resistor is 20 mA. [3 marks]
   c) Calculate the potential difference of the cell. [1 mark]

8 Figure 16.58 shows a circuit which includes a fixed resistor of 750 Ω and a component, X. The resistance of X changes with temperature. Figure 16.59 shows how the resistance of X changes with temperature.

   a) What is component X? [1 mark]
   b) At what temperature does X have a resistance of 250 Ω? [1 mark]
   c) Calculate the current flowing through the circuit when X has a resistance of 250 Ω. [3 marks]

The component is now placed in a beaker of water and warmed from 20°C to 100°C.
Practice questions

d) Describe how the reading on the voltmeter changes as the water warms from 20°C to 100°C. Give reasons for your answer. [3 marks]

![Figure 16.58]

9 a) The table shows the current in three different electrical appliances when connected to a 230 V a.c. supply.

<table>
<thead>
<tr>
<th>Appliance</th>
<th>Current in A</th>
</tr>
</thead>
<tbody>
<tr>
<td>Kettle</td>
<td>11.5</td>
</tr>
<tr>
<td>Bulb</td>
<td>0.05</td>
</tr>
<tr>
<td>Toaster</td>
<td>4.2</td>
</tr>
</tbody>
</table>

i) Which appliance has the greatest resistance? How does the data show this? [2 marks]

ii) The bulb is connected to the mains supply using a thin, twin-cored cable, consisting of live and neutral connections. State two reasons why this cable should not be used to connect the kettle to the mains supply. [2 marks]

b) Calculate the power rating of the kettle when it is operated from the 230 V a.c. mains supply. [3 marks]

The kettle is taken to the USA where the mains supply has a potential difference of 115 V.

c) i) Calculate the current flowing through the kettle when it is connected to a 115 V mains supply. [3 marks]

ii) The kettle is filled with water. The water takes 90 s to boil when working from the 230 V supply. Calculate how the time it takes to boil changes when the kettle operates on the 115 V supply. [3 marks]

10 a) In which circuit does the smallest current flow from the battery? [1 mark]

b) In which circuit does the largest current flow from the battery? [1 mark]

![Figure 16.60]
Working scientifically: Units and calibration

International System of Units (SI)
The three thermometers in Figure 16.61 are all measuring the same temperature but each one gives a different reading. This is because each thermometer has been calibrated using a different scale of units.

Scientists around the world have an agreed set of units that are used to measure quantities such as mass, time, current and temperature. These are known as the SI units. The units used in this book are SI units.

1 The table lists three quantities and the SI unit for that quantity.

Table 16.3

<table>
<thead>
<tr>
<th>Quantity</th>
<th>Unit</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Current</td>
<td>ampere</td>
<td>A</td>
</tr>
<tr>
<td>Resistance</td>
<td>ohm</td>
<td>Ω</td>
</tr>
<tr>
<td>Energy</td>
<td>joule</td>
<td>J</td>
</tr>
</tbody>
</table>

Draw your own table and list all of the quantities and their SI units found in this chapter.

2 Suggest why it is important that scientists around the world measure quantities using the same system of units.

Powers of ten
Some measurements that we make may be very small or very large. An electric current may be very small because the resistance of the circuit is very large. In this case, we may measure the current in milliamps and the resistance in kilohms.

The table lists the prefixes and powers of ten that you need to be able to use.

Table 16.4

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Power of ten</th>
</tr>
</thead>
<tbody>
<tr>
<td>tera</td>
<td>T</td>
<td>$10^{12}$</td>
</tr>
<tr>
<td>giga</td>
<td>G</td>
<td>$10^9$</td>
</tr>
<tr>
<td>mega</td>
<td>M</td>
<td>$10^6$</td>
</tr>
<tr>
<td>kilo</td>
<td>k</td>
<td>$10^3$</td>
</tr>
<tr>
<td>centi</td>
<td>c</td>
<td>$10^{-2}$</td>
</tr>
<tr>
<td>milli</td>
<td>m</td>
<td>$10^{-3}$</td>
</tr>
<tr>
<td>micro</td>
<td>μ</td>
<td>$10^{-6}$</td>
</tr>
<tr>
<td>nano</td>
<td>n</td>
<td>$10^{-9}$</td>
</tr>
</tbody>
</table>
Calibrating a voltmeter to measure temperature

Zach has designed the circuit shown in Figure 16.62 to measure temperature. The circuit includes a thermistor. We know that the resistance of a thermistor changes with temperature, so the thermistor is being used as the temperature sensor.

The reading on the voltmeter changes as the resistance of the thermistor changes. This means that the voltmeter can be used to measure temperature, but first it must be calibrated.

To calibrate the voltmeter to measure temperature Zach put the thermistor into a beaker of ice cold water at exactly 0°C. Zach then heated the water. Using an accurate thermometer, Zach measured and recorded different water temperatures and the reading of the voltmeter at those temperatures. Zach then drew the calibration graph shown in Figure 16.63.

1. Explain how Zach can use the calibration graph to convert a voltmeter reading into a temperature value.

2. Explain why Zach could not have drawn an accurate calibration graph if he had only put the thermistor into melting ice and boiling water.

3. Estimate the reading on the voltmeter if the thermistor were to be placed inside an oven at 120°C. (The rest of the circuit is outside the oven). To do this you must assume the pattern shown in Figure 16.63 continues – this is called extrapolating the results. Extrapolating results is easiest if the pattern is a straight line.

4. Explain how the resolution of the voltmeter as a thermometer changes as the temperature of the thermistor increases above 60°C.

**KEY TERM**

**Extrapolation** To make an estimate (or prediction) assuming that an existing trend or pattern continues to apply in an unknown situation.
In the 19th century scientists proposed the existence of atoms to explain some of their observations – for example the random motion of small particles of pollen, seen floating on water. By the early 20th century, we had worked out the size of atoms, but in the 21st century we can use a scanning tunnelling microscope to form an image of atoms in solids. In the photograph, each green sphere is a silicon atom. The image is generated by a computer, based on the tiny electron current detected by a probe near the surface of the silicon sample.

This chapter covers specification points 6.3.1.1 to 6.3.3.1 and is called Particle model of matter. It covers changes of state and the particle model, internal energy and energy transfers, and particle model and pressure.
Previously you could have learnt:

› Matter is made up of atoms and molecules.
› There are three states of matter: solid, liquid, gas.
› In a solid, atoms are packed together in regular patterns.
› In a liquid, atoms are in contact and are able to flow past each other.
› In a gas, atoms or molecules are spread out and are free to move around.
› The density of a substance measures the mass of that substance in a given volume. Solids and liquids are usually denser than gases.

Test yourself on prior knowledge

1 Name the three states of matter.
2 Explain why 1 m³ of gas (at atmospheric pressure) has less mass than 1 m³ of a solid material.
3 Why does 1 kg of water have a smaller volume than 1 kg of steam?

Density

A tree with a mass of 1000 kg obviously has a greater mass than a steel nail with a mass of 0.01 kg. However, sometimes we hear people say: ‘steel is heavier than wood.’ What they mean is that a piece of steel has a greater mass than a piece of wood with the same volume. Steel has a greater density than wood.

The density of a material is defined by the equation:

\[ \rho = \frac{m}{V} \]

where density, \( \rho \), is in kilograms per metre cubed, kg/m³
mass, \( m \), is in kilograms, kg
volume, \( V \), is in metres cubed, m³.

Example

Aluminium has a density of 2700 kg/m³. Calculate the volume of 135 g of aluminium.

Answer

\[ \rho = \frac{m}{V} \]

\[ 2700 = \frac{0.135}{V} \]

\[ V = \frac{0.135}{2700} \]

\[ V = 0.00005 \text{ m}^3 \]

\[ = 5 \times 10^{-5} \text{ m}^3 \]

Remember: always make sure to work in kg and m³.
An investigation to calculate the density of liquids and solids.

To calculate the density of a substance, the mass and volume of the substance must be measured. Density can then be calculated using the equation:

\[
density = \frac{\text{mass}}{\text{volume}}
\]

The density of a liquid

Method

1. Draw a table to write your results in. Make sure each column in the table has a heading that includes the quantity and the unit.
2. Use an electronic balance to measure the mass of an empty 100 millilitre (ml) measuring cylinder.
3. Take the measuring cylinder off the balance then carefully pour 20 ml of water into the measuring cylinder.
4. Measure the new mass of the measuring cylinder and water, then calculate the mass of water in the measuring cylinder.
5. Add another 20 ml of water to the measuring cylinder then measure the new mass. Repeat this by adding another 20 ml of water to the measuring cylinder.
6. You now have three sets of results for different masses and volume of water.

Analysing the results

1. Use each set of results to calculate a value for the density of water. Calculate the mean (average) value.
   - Note – a volume of 1 ml is the same as 1 centimetre cubed (cm³).
2. If you have mass in grams (g) and volume in cm³ then density will be in g/cm³. Change the density to kg/m³ by multiplying by 1000.

Taking it further

Olive oil and vinegar are often used to make a salad dressing. When left in a bottle they eventually separate out. The one with the highest density will sink to the bottom. Which one would you predict would go to the bottom of the bottle, the oil or the vinegar?

Measure the densities of olive oil and of vinegar and see if your prediction was right.

Questions

1. Why is it better to measure a large volume of liquid rather than a small volume?
2. What is meant by a resolution of 0.1 g?
The density of a regular solid

(a) The volume of a cuboid = length × width × height.
(b) The volume of a cylinder = \( \pi r^2 h \).

Method

1. Measure the dimensions of the solid. If it is a cuboid this will be the length, width and height. Measure each dimension several times in different places. If the measurements for, say, the length are different, calculate a mean value. Then calculate the volume of the solid.
2. Measure the mass of the solid.

Analysing the results

1. Calculate the density of the solid.
2. If you know the type of material your solid is, look up its true density. Calculate the difference between your value and the true value. Do you think your value is accurate?

Taking it further

Describe how to measure the density of a sheet of aluminium cooking foil.

Questions

1. Measuring the length of a cuboid three times may give three slightly different values. Suggest why.
2. Describe how you can measure the thickness of paper using an ordinary 30 cm ruler.

The density of an irregularly shaped solid

Method

1. Make sure the object to be used fits easily into the measuring cylinder.
2. Measure the mass of the object.
3. Put enough water into the measuring cylinder to submerge the object. Measure the volume of water in the measuring cylinder.
4. Tilt the measuring cylinder and slide the object in.
5. Measure the new position of the water surface in the measuring cylinder. Then calculate the volume of the object.

Analysing the results

1. Calculate the density of the object.

Taking it further

Use several different shaped objects all of the same material. This could simply be five or six stones. Measure the mass and volume of each object. Plot a graph of mass against volume. Your graph may look like Figure 17.4.
Use the graph to calculate the density of the objects. Your graph may include an anomalous data point. This could be due to a big measurement error but more likely one of the objects has a different density. If so then the object must be a different material.

Questions
1. Why should a graph of mass against volume go through the origin?
2. How does plotting a graph allow you to identify anomalous data?

KEY TERM
Anomalous result One that does not fit the expected pattern.

Figure 17.4 If the line misses the origin, there may have been an error in the measurements.

Test yourself

1. When answering an exam question a student wrote the following.
   'A cork floats on water because it is lighter than water. A stone sinks because it is too heavy to float on water.'
   Correct the mistakes in the student’s answer.

2. A student wants to determine the density of a cuboid of material. He takes these measurements.
   • Mass of the cuboid = 173.2 g
   • Length = 10.1 cm; Width = 4.8 cm; Height = 1.3 cm
   Calculate the density of the material in kg/m$^3$.

3. A geologist needs to determine the density of a rock. First she weighs the rock and calculates that its mass is 90 g. Next she measures the volume of the rock by immersing it in water.
   a) i) Use the information in Figure 17.5 to determine the volume of the rock in ml.
      ii) Express this volume in m$^3$. [$1$ ml = $10^{-6}$ m$^3$]
   b) Calculate the density of the rock in kg/m$^3$.

4. Copy the table and fill in the gaps.

<table>
<thead>
<tr>
<th>Material</th>
<th>Volume in m$^3$</th>
<th>Mass in kg</th>
<th>Density in kg/m$^3$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>3</td>
<td>3000</td>
<td></td>
</tr>
<tr>
<td>Alcohol</td>
<td>3200</td>
<td>800</td>
<td></td>
</tr>
<tr>
<td>Titanium</td>
<td>0.5</td>
<td>0.2</td>
<td>4500</td>
</tr>
<tr>
<td>Cork</td>
<td>0.02</td>
<td>390</td>
<td></td>
</tr>
</tbody>
</table>

Figure 17.5 Measuring the volume of an irregular solid.

Show you can... Show that you understand the definition of density by describing an experiment to calculate the density of a liquid. Explain what measurements you will take and the calculations you will do.
Solid, liquids and gases

Ice, water and steam are three different states of the same substance. We call these three states solid, liquid and gas.

- **Solid**: In a solid the atoms (or molecules) are packed in a regular structure. The atoms cannot move out of their fixed position, but they can vibrate. The atoms are held close together by strong forces. So it is difficult to change the shape of a solid.

- **Liquid**: The atoms (or molecules) in a liquid are also close together. The forces between the atoms keep them in contact, but atoms can move from one place to another. A liquid can flow and change shape to fit any container. Because the atoms are close together, it is very difficult to compress a liquid.

- **Gas**: In a gas the atoms (or molecules) are separated by relatively large distances. The forces between the atoms are very small. The atoms in a gas are in a constant random motion. A gas can expand to fill any volume.

<table>
<thead>
<tr>
<th>Material</th>
<th>Density in kg/m$^3$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lead (solid)</td>
<td>11 400</td>
</tr>
<tr>
<td>Glass (solid)</td>
<td>2 500</td>
</tr>
<tr>
<td>Water (liquid)</td>
<td>1 000</td>
</tr>
<tr>
<td>Lithium (solid)</td>
<td>500</td>
</tr>
<tr>
<td>Cork (solid)</td>
<td>200</td>
</tr>
<tr>
<td>Air (gas)</td>
<td>1.3</td>
</tr>
<tr>
<td>Hydrogen (gas)</td>
<td>0.09</td>
</tr>
</tbody>
</table>

- Lead has a much higher density than lithium because the atoms of lead have a much greater mass than lithium atoms. (The lead atoms are only slightly larger than lithium atoms.)

- The density of air is greater than the density of hydrogen, because the nitrogen and oxygen molecules in the air have a greater mass than hydrogen molecules.
Internal energy

Energy is stored inside a system by the particles (atoms or molecules) that make up that system. This is called internal energy. Internal energy is the total kinetic and potential energy of all the particles that make up the system.

We can use a model of several balls and springs to help us understand the nature of internal energy in a solid. The balls represent the atoms and the springs represent the forces or ‘bonds’ that keep the atom in place.

In Figure 17.7a) the middle atom has been displaced to the right. Now potential energy is stored in the stretched bond. In Figure 17.7b) the atom has kinetic energy as it moves to the left.

Heating

Heating changes the energy stored within a system by increasing the energy of the particles that make up the system.

- Heating can increase the temperature of a system. For example, when a gas is heated, the atoms (or molecules) move faster and the kinetic energy of the atoms increases.
- Heating a system can also cause a change of state; for example, when a solid melts to become a liquid. Usually when a solid melts, there is a small increase in volume, as the solid turns to liquid. The atoms increase their separation and there is an increase in the potential energy stored. So the internal energy increases.

Changes of state

We use these terms to describe changes of state.

- Melting occurs when a solid turns to a liquid. The internal energy of the system increases.
- Freezing occurs when a liquid turns to a solid. The internal energy of the system decreases.
- Boiling or evaporation occurs when a liquid turns to a gas. The internal energy of the system increases.
- Condensation occurs when a gas turns to a liquid. The internal energy of the system decreases.
- Sublimation occurs when a solid turns directly into a gas. The internal energy of the system increases. Sublimation is rare. An example is carbon dioxide (CO\textsubscript{2}): solid CO\textsubscript{2} (dry ice) turns directly into the gas CO\textsubscript{2} missing out the liquid state at normal atmospheric pressure.
Specific heat capacity

A change of state of a substance is a physical change. The change does not produce a new substance and the process can be reversed. For example, a cube of ice from the freezer can be allowed to melt into water. The water can be put back into its container and then into the freezer. The water will freeze back into ice. No matter what its state, water or ice, the mass is the same. So, when a substance changes state, the mass is conserved. This is because the total number of particles (atoms or molecules) stays the same.

Test yourself

7 a) What is meant by a change of state of a substance?  
b) Give two examples of changes of state.
8 Which of the following changes are physical changes?  
   - Melting snow  
   - Burning a matchstick  
   - Breaking a matchstick  
   - Boiling an egg  
   - Mixing salt and sugar together.
9 a) What happens to the internal energy of a system when the system is heated?  
b) How is it possible to heat a system without the temperature of the system increasing?

Specific heat capacity

When the temperature of a system is increased by supplying energy to it, the increase in temperature depends on:

- the mass of the substance heated
- the type of material (what the substance is made of)
- the energy put into the system.

Water needs much more energy to increase its temperature by 1 °C than the same mass of concrete. This also means that when water cools by 1 °C, it gives out more energy than the same mass of concrete cooling by 1 °C.

▶ Figure 17.9 a) 4200 joules of energy are needed to increase the temperature of 1 kg of water by 1 °C.  
b) 800 joules of energy are needed to increase the temperature of 1 kg of concrete by 1 °C.
Based on Figure 17.9, we say that the specific heat capacity of water is 4200 joules per kilogram per degree Celsius (J/kg °C).

The specific heat capacity of a substance is the amount of energy required to raise the temperature of one kilogram of the substance by one degree Celsius.

To calculate the change in thermal energy in a substance we use the equation:

\[ \Delta E = mc \Delta \theta \]

Where change in thermal energy is in joules, \( J \)

mass is in kilograms, \( kg \)

specific heat capacity is in joules per kilogram per degree Celsius, \( J/kg °C \)

temperature change is in degrees Celsius, \( °C \).

Table 17.3 gives some examples of specific heat capacities for various substances at 20 °C.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific heat capacity J/kg °C</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>4200</td>
</tr>
<tr>
<td>Alcohol</td>
<td>2400</td>
</tr>
<tr>
<td>Ice</td>
<td>2100</td>
</tr>
<tr>
<td>Dry air</td>
<td>1000</td>
</tr>
<tr>
<td>Aluminium</td>
<td>880</td>
</tr>
<tr>
<td>Concrete</td>
<td>800</td>
</tr>
<tr>
<td>Glass</td>
<td>630</td>
</tr>
<tr>
<td>Steel</td>
<td>450</td>
</tr>
<tr>
<td>Copper</td>
<td>380</td>
</tr>
<tr>
<td>Lead</td>
<td>160</td>
</tr>
</tbody>
</table>

The specific heat capacity of water

Water has a very high specific heat capacity. This means that 1 kg of water requires a lot of energy to heat it up and a lot of energy must be transferred from the water when it cools down. This high specific heat capacity is very important.

- We are made mostly of water. A high specific heat capacity of water means that our body temperature does not increase too much when we take exercise or cool too quickly when we go outside on a cold day.
- Water is used for keeping many homes warm. A house central heating system pumps hot water around the house. The hot water transfers a lot of energy as it flows through radiators. If water had a low specific heat capacity, water would cool down before it got to some of the radiators in your house.
Latent heat

When you heat a pan of water on the cooker, energy is transferred to the water and the temperature of the water increases. After a while the water begins to boil, and the temperature of the water stays constant at 100 °C. Yet, the cooker is still supplying energy to the water at the same rate. So where is the energy transferred to now? The answer is that energy is transferred into the internal energy of the steam. The molecules in steam at 100 °C have more internal energy than the same molecules of water at 100 °C.

The energy needed for 1 kg of a substance to change state is called specific latent heat.

The specific latent heat of a substance is the amount of energy required to change the state of one kilogram of the substance with no change in temperature.

The energy required to change the state of a substance can be calculated using this equation:

\[ E = m L \]

energy required for a change of state = mass \times specific latent heat

where energy is in joules, J
mass is in kilograms, kg
specific latent heat is in joules per kilogram, J/kg.
Matter has three states: solid, liquid and gas. So a substance has two specific latent heats.

- The **specific latent heat of fusion** is the energy required to change 1 kg of a solid into 1 kg of a liquid at the same temperature.
- The **specific latent heat of vaporisation** is the energy required to change 1 kg of a liquid into 1 kg of a gas (vapour) at the same temperature.

### Melting and freezing

When a substance melts, energy must be supplied to the substance.

When a substance freezes (or solidifies), energy is transferred from the substance to the surroundings.

### Vaporising and condensing

When a substance vaporises, energy is supplied to the substance to turn it from a liquid into a gas.

When a substance condenses, energy is transferred from the substance as it changes from a gas into a liquid.

## Measuring latent heat

Figure 17.12 shows how you can calculate the latent heat of vaporisation of water.

- A beaker of water is placed on top of a balance. The beaker is on a heatproof mat. The water is then brought to boiling point with a heater. At the moment the water boils, the joulemeter is reset to zero.
- The heater is then allowed to boil the water for 5 minutes.
- The following measurements are taken:
  1. Joulemeter reading after 5 minutes – 60 kJ
  2. Mass of beaker and contents at the start – 968 g
  3. Mass of beaker and contents after 5 minutes – 944 g.

### Calculation

**Mass of water turned to steam** = 968 g – 944 g

= 24 g = 0.024 kg

\[ E = mL \]

So 60 000 \( = 0.024 \times L \)

\[ L = \frac{60\,000}{0.024} \]

= 2 500 000 J/kg

= 2.5 MJ/kg
Cooling graphs

When a boiling tube containing water is heated and then left to cool down, the temperature of the water drops gradually (see Figure 17.13).

When the temperature of the water is high, the temperature drops quickly. When water is closer to room temperature, the temperature drops more slowly.

If a substance changes state as it cools, the cooling curve takes a different shape. Ethanamide is a substance that melts at 80 °C. When a boiling tube containing ethanamide (Figure 17.14) is allowed to cool from 100 °C, it cools quickly from 100 °C to 80 °C (Figure 17.15). Then the temperature remains constant for a few minutes as the ethanamide solidifies (or freezes). Although the boiling tube continues to transfer energy to the surroundings, the temperature of the ethanamide remains constant at 80 °C. This is possible because the ethanamide releases energy as the internal energy of its molecules decreases. When all the ethanamide has solidified, its temperature begins to fall again.

Test yourself

13 a) When we take exercise we sweat. The sweat evaporates from our skin (Figure 17.16). Why does sweating help us stay cool?
   b) Even on a warm day having wet skin can soon make you feel cold (Figure 17.17). Explain why.

14 A solid is heated at a constant rate until it becomes a gas. Figure 17.18 shows how the temperature of the substance increases with time.
17 Particle model of matter

The particle model of gases

As a result of studying the behaviour of gases, we have built up a model (or theory) to help us understand, explain and predict the properties of gases. This is called the particle model or kinetic theory of gases. The main points of the model are listed here.

- The particles in a gas (molecules or atoms) are in a constant state of random motion.
- The particles in a gas collide with each other and the walls of their container without losing any of their kinetic energy.
- The temperature of the gas is related to the average kinetic energy of the molecules.
- As the kinetic energy of the molecules increases the temperature of the gas increases.

Gas pressure

When the particles of a gas collide with a wall of their container, the particles exert a force on the wall. Figure 17.20 shows three particles bouncing off a container wall. Each particle exerts a force on the wall at right angles to the wall.
The pressure inside a container of gas, with a fixed volume, is increased when the temperature of the gas is increased. When the temperature of a gas increases, the average speed of the particles in the gas increases. This means that the particles hit the walls of the container with a greater force and the particles hit the walls more frequently. So the pressure increases.

**Demonstrating gas pressure**

Your teacher might demonstrate the effect of gas pressure as follows:

- Take a tin with a close fitting lid and put a small amount of water in it. Press the lid firmly in place.
- Then put the tin on a tripod and heat with a Bunsen burner. After a while the lid flies off.

So why does the lid fly off?

- In Figure 17.21a) the molecules inside the tin move at the same speed as the molecules outside the tin. There is no force on the tin lid.
- In Figure 17.21b) two things have happened. As the tin warms up, some water evaporates so the number of molecules of gas inside the tin increases. Then the molecules travel faster as the temperature rises (shown with longer arrows on the molecules). The molecules inside the tin exert a force on the lid large enough to blow it off.

▲ Figure 17.21 Safety note: This experiment should be done behind a safety screen and everyone should wear safety glasses.
Chapter review questions

1 The sides of a block of wood measure 4.0 cm, 3.0 cm and 5.0 cm. The block of wood has a mass of 30.0 grams.
   a) Calculate the volume of the wood in m³.
   b) Calculate the density of the wood in kg/m³.

2 A man buys a 'gold' ornament from an antiques shop. He decides to check if the ornament is made of solid gold.
   The results of his measurements are shown below.
   • mass of ornament = 320 g
   • volume of ornament = $26 \times 10^{-6}$ m³
   a) Explain how the man might have measured the mass and volume of the ornament.
   b) Calculate the density of the ornament.
   c) Use the data below to suggest what the man might find if he cuts his ornament in half.
      • density of gold = 19 300 kg/m³
      • density of lead = 11 600 kg/m³.

3 When a drop of ether is placed on the skin, the skin feels cold. Explain why.

4 a) State one difference between the arrangement of the molecules in water and the molecules in ice. Draw a diagram to illustrate your answer.
   b) An ice cube at a temperature of 0 °C is more effective in cooling a drink that the same mass of water at 0 °C. Give the reason why.
   c) Give the reason why ice floats on water.

5 Use the information in Table 17.1 to help you answer these questions.
   a) A heater supplies 200 kJ of energy to 40 kg of dry air at a temperature of 15 °C. Calculate the temperature of the air after it has been heated.
   b) A block of concrete has a mass of 60 kg. It cools down from 48 °C to 13 °C. Calculate the energy transferred from the block.

6 a) Explain in terms of molecular motion why a gas exerts a pressure on the walls of its container.
   b) The pressure in a container of gas increases when the temperature increases. Explain why.

7 A pan of water is placed on top of a cooker. The cooker transfers energy to the water at a rate of 500W.
   a) When the water is boiling, the pan is left on the cooker for 5 minutes. Calculate the energy transferred to the cooker in this time.
   b) Calculate the mass of water that turns into steam in 5 minutes. (The specific latent heat of vaporisation of water is 2.5 MJ/kg.)
Practice questions

1 Which of the following is the correct unit for density?  
   kg/m² kg/m kg/m³ m³/kg [1 mark]

2 Figure 17.22 represents four measuring cylinders each containing a liquid. The mass and volume of the liquid in each cylinder are given.

<table>
<thead>
<tr>
<th>Mass (g)</th>
<th>Volume (cm³)</th>
</tr>
</thead>
<tbody>
<tr>
<td>W</td>
<td>80</td>
</tr>
<tr>
<td>X</td>
<td>100</td>
</tr>
<tr>
<td>Y</td>
<td>100</td>
</tr>
<tr>
<td>Z</td>
<td>80</td>
</tr>
</tbody>
</table>

   ▲ Figure 17.22

Which two measuring cylinders could contain the same liquid?  
   a) W and X  
   b) W and Y  
   c) X and Y  
   d) X and Z [1 mark]

3 Each of the following statements describes either a solid, a liquid or a gas.

   a) It takes the shape of its container, but does not always fill the container. [1 mark]  
   b) The particles are in a regular pattern. [1 mark]  
   c) The particles are free to move over each other. [1 mark]  
   d) The particles move in random directions. [1 mark]  
   e) The particles always fill the whole of their container. [1 mark]  
   f) The particles vibrate about a fixed position. [1 mark]

4 Energy is supplied to a substance and its temperature remains the same. Explain how this is possible. [2 marks]

5 Describe the differences between the arrangement of the atoms in a solid and in a gas. [4 marks]

6 Explain how you would use a measuring cylinder, electronic balance and some glass marbles to calculate the density of glass. [4 marks]

7 The apparatus in Figure 17.23 is used to heat up a block of metal of mass 2 kg. When the heater is turned on, the temperature of the block of metal increases as shown in Figure 17.24.

   a) Use the graph to determine the temperature rise of the metal in the first 10 minutes of heating. [1 mark]

   b) During 10 minutes of heating, 48,000 J of energy is supplied to the block. Calculate the specific heat capacity of the block. [4 marks]

   c) Use the information in part (b) and information from the graph to show that the power of the heater is 80 W. [3 marks]

8 Figure 17.25 shows a heater at the bottom of a boiling tube of solid wax. The heater is then connected to a power supply. A joulemeter measures the energy supplied to the heater as it melts the wax. The graph in Figure 17.26 shows how the temperature of the wax changes with the energy supplied.
a) State the melting temperature of wax. [1 mark]

b) The temperature of the wax remains constant as the wax melts. Explain why. [2 marks]

c) The mass of the wax is 50 g. Use information from the graph to calculate the specific latent heat of fusion for the wax. Give the unit. [4 marks]
We all know what a scale model is and you have probably seen a model aircraft. It may not have been an exact replica but you would be able to recognise it as an aircraft. If you can’t, then the model probably needs replacing.

In science we often use a physical model to help visualise objects or systems that are too big, too small or impossible to see. A scale model of the Solar System helps us to understand the position and distances between the planets and the Sun.

In this chapter you have read about the particle model of a gas. Atoms and molecules are too small for us to see directly so a physical model can help us to visualise what is happening and understand the real thing.

1 Describe another physical model given in this chapter.

Scientific models

A scientific model is an idea or related group of ideas used to explain something in the real world. A model is sometimes called a theory. Some people think that the word ‘theory’ means a guess; it does not. A theory or model is the idea used to explain observations and patterns in data.

A scientific model should be able to:

► explain observations
► be used to predict outcomes
► fit with other ideas in science.

Scientists use graphs, diagrams, equations, computer graphics and physical structures to represent and communicate models to others.

The particle model of a gas can be represented by mathematical equations. The equations can be used to explain and predict how changing one variable,
17 Particle model of matter

for example the volume, affects the pressure and temperature of the gas.

2 The equation $F = ma$ is a mathematical model. What would you predict happens to the acceleration of an object when the resultant force on the object is increased? How would you be able to test this model?

Why do models change?

A scientific model is only as good as the evidence that supports it. A model is supported if any predictions made using the model turn out to be correct. However, a test that gives data which the model cannot predict provides evidence that the model may be wrong.

So models change to give a better fit to any new experimental results or observations.

In Chapter 18 you can read how new experimental evidence led to the plum pudding model of the atom being replaced by the nuclear model. This model itself has been changed several times, each time to explain the most up-to-date observations.

Alternative models

Sometimes scientists have more than one model to explain the same thing. Light is sometimes described as a wave. This is one model. However, sometimes light is better explained using the idea that it is a stream of particles. So is one model right and the other wrong? The answer is no. Each of the models is appropriate but in different situations.

Scientists know that they cannot explain everything. Sometimes there are alternative models but insufficient evidence to support or reject any of them. For example, how will the Universe end? Will it expand for ever, will it eventually shrink or will it stop expanding and stay at that size? A group of scientists using a new mathematical model believe the Universe will rip itself apart. However, there is no need to panic. The model predicts it will not happen for another 22 billion years.

3 What would be needed for one of the models explaining the end of the Universe to be accepted by scientists and the other models rejected?
Henry Becquerel discovered radioactivity in 1896. Over the last century, we have learnt how to use radioactive materials safely and to put them to good use. Radioactive sources are used in industry, agriculture and in medicine. The image shows the concentration of radioactive sugar 2 hours after tracer molecules were fed into a plant. The red colour shows a high concentration of sugar in the young leaves of the plant, which are growing. The fast growing young leaves take the sugar from the older leaves, which appear blue.

This chapter covers specification points 6.4.1.1 to 6.4.2.4 and is called Atomic structure. It covers atoms and isotopes as well as atoms and nuclear radiation.
Previously you could have learnt:

- All materials are made up of tiny particles called atoms.
- Elements are made up of only one type of atom.
- An atom has a very small positively charged nucleus.
- The nucleus contains protons and neutrons.
- Negatively charged electrons orbit the nucleus.
- The proton carries a positive charge and the electron carries a negative charge of the same size as the proton.
- An atom is neutral in charge, because the positive charge on the nucleus is balanced by the negative charge of the electrons.

Test yourself on prior knowledge

1. Which two particles are in the nucleus of an atom?
2. Which particle is more massive, the proton or the electron?
3. Why are the number of protons and electrons the same for a particular atom?

Atoms and isotopes

(a) hydrogen atom (b) lithium atom

Figure 18.1 These figures show the numbers of protons, neutrons and electrons in a) a hydrogen and b) a lithium atom.

**KEY TERMS**

- **Proton** A positively charged particle found in the nucleus of an atom.
- **Neutron** A neutral particle found in the nucleus of an atom.
- **Electron** A negatively charged particle that orbits the nucleus of an atom.

**Neutrons, protons and electrons**

We cannot see atoms directly because they are so small. However, indirect measurements show us that the radius of an atom is about $10^{-10}$ m (0.000 000 000 1 m). The radius of the nucleus is less than 1/10 000 of the radius of the entire atom.

Inside the nucleus there are two types of particle, protons and neutrons. The protons and neutrons have approximately the same mass. A proton has a positive charge and the neutron has no charge. Outside the nucleus there are electrons which orbit the nucleus at distances of about $10^{-10}$ m. Electrons are able to move in different orbits around the nucleus, and change their orbit by absorbing or emitting electromagnetic radiation.

Electrons have very little mass in comparison with protons and neutrons. Electrons carry a negative charge which is the same size as the positive charge on the proton. Because protons and neutrons are much more massive than electrons, most of the mass of an atom is in its nucleus.
Atoms and ions

A hydrogen atom has one proton and one electron; it is electrically neutral because the charges of the electron and proton cancel each other out. A helium atom has two protons and two neutrons in its nucleus, and two electrons outside that. The helium atom is also neutral because it has the same number of electrons as it has protons.

Ions

Atoms are electrically neutral because the number of protons balances exactly the number of electrons. However, it is possible either to add extra electrons to an atom or to take them away. When an electron is added to an atom a negative ion is formed; when an electron is removed a positive ion is formed. Ions are made in pairs because an electron that is removed from one atom attaches itself to another atom, so a positive and negative ion pair is formed.

Atomic and mass numbers

The number of protons in the nucleus of an atom determines what element it is. Hydrogen atoms have one proton, helium atoms two protons, uranium atoms 92 protons. The number of protons in the nucleus is the same as the number of electrons surrounding the nucleus. The number of protons in the nucleus is called the atomic number of the atom (symbol Z). So the proton number of hydrogen is 1; \( Z = 1 \).

The mass of an atom is determined by the number of neutrons and protons added together. Scientists call this number the mass number of an atom.

\[
\text{atomic number} = \text{number of protons} \\
\text{mass number} = \text{number of protons plus neutrons}
\]

For example, an atom of sodium has eleven protons and twelve neutrons. So its atomic number is 11 and its mass number is 23. To save time in describing we can write it as \( ^{23}\text{Na} \); the mass number appears on the left and above the symbol Na, for sodium, and the atomic number on the left and below.
Isotopes

Not all the atoms of a particular element have the same mass. For example, two sodium atoms might have mass numbers of 23 and 24. The nucleus of each atom has the same number of protons, 11, but one atom has 12 neutrons and the other 13 neutrons. Atoms of the same element (sodium in this case) that have different masses are called isotopes. These two isotopes of sodium can be written as sodium-23, $^{23}_{11}Na$, and sodium-24, $^{24}_{11}Na$.

Tip

The symbols $^{23}_{11}Na$ describe only the nucleus of a sodium atom.

Key Term

Isotopes: Different forms of a particular element. Isotopes have the same number of protons but different numbers of neutrons.

Test yourself

1. a) What is the approximate radius of an atom? 
   
   $10^{-3}$ m  $10^{-6}$ m  $10^{-10}$ m

   b) Use an answer from the box to complete the sentence below.

   $1000$ $10000$ $100000$

   The radius of an atom is about ____________ times larger than the radius of a nucleus.

2. The diagram shows the nuclei of three atoms. Which two atoms are isotopes of the same element? Give a reason for your answer.

   A  B  C

   Figure 18.2

3. a) A nitrogen atom has 7 protons, 7 neutrons and 7 electrons.

   i) What is the atomic number of nitrogen? 

   $7$  $14$  $21$

   Give a reason for your answer.

   ii) What is the mass number of nitrogen? 

   $7$  $14$  $21$

   Give a reason for your answer.

   b) Explain why a nitrogen atom is neutral.

4. Calculate the number of protons and neutrons in each of the following nuclei.

   a) $^{17}_{8}O$  b) $^{200}_{80}Hg$  c) $^{238}_{92}U$  d) $^{3}_{1}H$

5. Gadolinium-156 and gadolinium-158 are two isotopes of gadolinium, which has an atomic number of 64.

   a) Explain what the numbers 64, 156 and 158 mean.

   b) i) What do these two isotopes have in common?

   ii) How are the two isotopes different?

6. Explain the term isotope.

7. The radius of a magnesium nucleus is $3.0 \times 10^{-15}$ m and the radius of a magnesium atom is $1.5 \times 10^{-10}$ m. How many times larger is the radius of a magnesium atom than the radius of a magnesium nucleus?

Show you can...

Complete this task to show that you understand the model of the atom.

Write a paragraph to explain and describe the structure of an atom. In your answer include these words: nucleus, proton, neutron, electron, mass number and atomic number.
Discovery of the nucleus

The discovery of the nucleus and the development of our understanding of the atom provide a clear example of how new experimental evidence can lead to a scientific model being changed or replaced.

Democritus, a Greek philosopher, lived from 460–370 BC. He was the first person to suggest the idea of atoms as small particles that cannot be cut or divided.

The discovery of the electron

In 1897, J. J. Thompson discovered that electrons were emitted from the surface of hot metals. Thompson showed that electrons are negatively charged and that they are much less massive than atoms. This discovery led to a change in the accepted atomic theory.

In 1904, Thompson proposed a new model for the atom. His idea was that atoms were made up of a ball of positive charge with electrons dotted around inside it (Figure 18.3). This idea is known as the ‘plum pudding’ model of the atom as it looks rather like a solid pudding with plums in it.

The nuclear model of the atom

In 1909, Geiger and Marsden discovered a way of exploring the insides of atoms. They directed a beam of alpha particles at a thin sheet of gold foil. Alpha particles were known to be positively charged helium ions, He\(^{2+}\), which travel very quickly. They had expected all of these energetic particles to pass straight through the thin foil because they thought the atom was like Professor Thompson’s soft plum pudding model. Much to their surprise they discovered that although most of them travelled through the foil without any noticeable change of direction, a very small number of the alpha particles bounced back.

Rutherford suggested that the deflection of an alpha particle was due to an electrostatic interaction between it and a very small charged nucleus. He also suggested that the nucleus must be massive, because
it did seem to be moved by the energetic alpha particle. In Rutherford’s original paper he suggested that the nucleus might be either positively or negatively charged. A positive nucleus would repel the positively charged alpha particle backwards, and a negative nucleus might pull the alpha particle round it, in the same way that a comet falling towards the Sun has its direction changed by the pull of gravity.

**Alpha scattering explained**

Now we know that the gap between the nucleus and electrons is large; the diameter of the atom is about 20 000 times larger than the diameter of the nucleus itself (Figure 18.6). After Rutherford’s original paper, scientists confirmed that the nucleus is positively charged, and that allows us to explain the scattering of the particles as follows. Because so much of the atom is empty space (Figure 18.6), most of the alpha particles could pass through it without getting close to the nucleus. Some particles pass close to the nucleus and so the positive charges of the alpha particle and the nucleus repel each other, causing a small deflection. A small number of particles met the nucleus head on; these are turned back the way they came (Figure 18.5). The fact that only a very tiny fraction of the alpha particles bounce backwards tells us that the nucleus is very small indeed. Rutherford proposed that all the mass and positive charge of an atom are contained in the nucleus and that the electrons outside the nucleus balance the charge of the protons.

Later experiments led to the idea that the positive charge of any nucleus could be subdivided into a whole number of smaller particles. These particles were called protons, and the positive charge on a proton was discovered to be the same size as the negative charge on an electron. Rutherford did not know that there are *neutrons* in the nucleus; these were discovered by James Chadwick in 1932. Chadwick’s discovery allowed scientists to account for the mass of the atom.

---

**The Bohr model of the atom**

In 1913, Niels Bohr suggested a model of the atom in which electrons move round the nucleus in *circular orbits at specific distances from the nucleus*. In this model electrons can change their orbits. Figure 18.7 shows the Bohr model of a hydrogen atom, with the first three energy levels.

The Bohr model was successful in explaining why hydrogen emits particular wavelengths of electromagnetic radiation. In Figure 18.7 an electron emits electromagnetic radiation and so loses some energy. It moves closer to the nucleus as it falls from level 3 to level 2. The reverse process is possible too: if an electron absorbs electromagnetic radiation it can move into a higher energy level further away from the nucleus.

Although the Bohr model was successful up to a point, it does not allow a full explanation of the behaviour of electrons in larger atoms. However, Bohr adapted his model for the hydrogen atom by suggesting that electrons in larger atoms are also confined to specific
Atoms and nuclear radiation

Henry Becquerel discovered radioactivity in 1896. He placed some uranium salts next to a photographic plate which had been sealed in a thick black bag to prevent light exposing the plate. When the plate was later developed it had been affected as if it had been exposed to light. Becquerel realised that new particles were emitted from uranium salts, which passed through the bag.

Nuclear decay

The nuclei of most atoms are very stable. The atoms that we are made of have been around for thousands of millions of years. Atoms may lose or gain a few electrons during chemical reactions, but the nucleus does not change during such processes.

However, there are some atoms that have unstable nuclei which throw out particles to make the nucleus more stable. This is a random process that depends only on the nature of the nucleus. The rate at which particles are emitted from a nucleus is not affected by other factors such as temperature or chemical reactions. One element discovered that emits these particles is radium, and the name radioactivity is given to this process.

Test yourself

8 a) How is the mass of the atom distributed in the plum pudding model?
   b) Where is most of the mass of the atom in the nuclear model?
9 Explain how you know that the polarity of the charge on an alpha particle is the same as that of the nucleus.
10 Why did most of the alpha particles pass through the foil without being deflected?
11 Describe the plum pudding model of the atom.
12 a) Describe Bohr’s model for the atom. Draw a diagram to help your explanation.
   b) Explain what happens to an electron when
      i) it absorbs electromagnetic radiation.
      ii) it emits electromagnetic radiation.
13 Figure 18.8 shows the path of an alpha particle being deflected by the charge of a gold nucleus.
   Make a copy of the diagram to show the paths of two more alpha particles B and C.

Show you can...

In the 19th century, scientists thought that atoms were the smallest particles. Explain what evidence led scientists to change this model of the atom.
There are four types of radioactive emission.

**Alpha particles** ($\alpha$) are identical to the nuclei of helium atoms. The alpha particle is formed from two protons and two neutrons, so it has a mass number of four and an atomic number of two. When an alpha particle is emitted from a nucleus it causes the nucleus to change into another nucleus with a mass number four less and an atomic number two less than the original one, for example:

\[
\begin{align*}
^{238}_{92}\text{U} & \rightarrow ^{234}_{90}\text{Th} + \alpha \\
\text{uranium nucleus} & \text{thorium nucleus alpha particle (helium nucleus)}
\end{align*}
\]

This is called **alpha decay**.

**Beta particles** ($\beta$) are fast moving electrons. In a nucleus there are only protons and neutrons, but a beta particle is made and ejected from a nucleus when a neutron turns into a proton and an electron. Since an electron has a very small mass, when it leaves a nucleus it does not alter the mass number of that nucleus. However, the electron carries away a negative charge so the removal of an electron increases the atomic number of a nucleus by 1. For example, carbon-14 decays into nitrogen by emitting a beta particle:

\[
\begin{align*}
^{14}_{6}\text{C} & \rightarrow ^{14}_{7}\text{N} + \beta \\
\text{carbon nucleus} & \text{nitrogen nucleus beta particle (electron)}
\end{align*}
\]

When some nuclei decay by sending out an alpha or beta particle, they also give out a **gamma ray**. Gamma ($\gamma$) rays are electromagnetic waves, like radio waves or visible light. They carry away from the nucleus a lot of energy, so that the nucleus is left in a more stable state. Gamma rays have no mass or charge, so when one is emitted there is no change to the mass or atomic number of a nucleus.

**Neutrons** are emitted from some highly unstable nuclei. The effect of this is to reduce the mass number by 1, but the atomic number does not change. Neutron emission is rare, but neutrons are a dangerous radiation. An example of neutron emission from helium-5 is given below:

\[
\begin{align*}
^{5}_{2}\text{He} & \rightarrow ^{4}_{2}\text{He} + n \\
\text{helium nucleus} & \text{helium nucleus neutron}
\end{align*}
\]

<table>
<thead>
<tr>
<th>Table 18.1 A summary of the four types of radioactive emission.</th>
</tr>
</thead>
<tbody>
<tr>
<td>Radiation emitted from nucleus</td>
</tr>
<tr>
<td>-------------------------------</td>
</tr>
<tr>
<td><strong>alpha ($\alpha$) decay</strong></td>
</tr>
<tr>
<td><strong>beta ($\beta$) decay</strong></td>
</tr>
<tr>
<td><strong>gamma ($\gamma$) decay</strong></td>
</tr>
<tr>
<td><strong>neutron (n) decay</strong></td>
</tr>
</tbody>
</table>
Ionisation

All types of radiation cause ionisation. This is why we must be careful when we handle radioactive materials. The radiation makes ions in our bodies and these ions can then damage our body tissues.

Your teacher can show the ionising effect of radium by holding some close to a charged gold leaf electroscope (Figure 18.11). The electroscope is charged positively at first so that the gold leaf is repelled from the metal stem. When a radium source is brought close to the electroscope, the leaf falls, showing that the electroscope has been discharged. The reason for this is that the alpha particles from the radium create ions in the air above the electroscope. This is because the charges on the alpha particles pull some electrons out of air molecules (Figure 18.10). Both negative and positive ions are made; the positive ones are repelled from the electroscope, but the negative ones are attracted so that the charge on the electroscope is neutralised. It is important that you understand that it is not the charge of the alpha particles that discharges the electroscope, but the ions that they produce.

Test yourself

14 What is the nature of each of the following:
   a) an alpha particle
   b) a beta particle
   c) a gamma ray?

15 Which row in the table, A, B or C, shows what happens to the mass number and atomic number of a nucleus when a gamma ray is emitted from it?

Table 18.2

<table>
<thead>
<tr>
<th>Mass number</th>
<th>Atomic number</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>Increases</td>
</tr>
<tr>
<td>B</td>
<td>Does not change</td>
</tr>
<tr>
<td>C</td>
<td>Decreases</td>
</tr>
<tr>
<td>A</td>
<td>Decreases</td>
</tr>
<tr>
<td>B</td>
<td>Does not change</td>
</tr>
<tr>
<td>C</td>
<td>Increases</td>
</tr>
</tbody>
</table>
Detecting particles

We make use of the ionising properties of alpha, beta and gamma radiations to detect them. Those three radiations are emitted by radioactive sources permitted in schools.

**Safety note:** Only your teacher can do this experiment as you are not allowed to handle radioactive sources until you are over 16.

Figure 18.12 shows how your teacher can use a Geiger-Müller (GM) tube and a radioactive source to investigate the range of different radiations.

1. A source emitting alpha particles, for example, can then be placed in front of the GM tube. By varying the separation (x) of the tube and the source, the range of the radiation may be calculated.

2. You can also check which materials stop a type of radiation. Now you keep the distance, x, constant. Then you can place various absorbers such as paper or metal foils in between the source and the GM tube.

Properties of radiation

Alpha particles travel about 5 cm through air but can be stopped by a sheet of paper (Figure 18.13). They ionise air very strongly.

Beta particles can travel several metres through air. They are stopped by a sheet of aluminium that is a few millimetres thick (Figure 18.13). They do not ionise air as strongly as alpha particles.
**Table 18.3** The range and penetration of radiation.

<table>
<thead>
<tr>
<th>Radiation</th>
<th>Nature</th>
<th>Range in air</th>
<th>Ionising power</th>
<th>Penetrating power</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alpha $\alpha$</td>
<td>Helium nucleus</td>
<td>A few centimetres</td>
<td>Very strong</td>
<td>Stopped by paper</td>
</tr>
<tr>
<td>Beta $\beta$</td>
<td>Electron</td>
<td>A few metres</td>
<td>Medium</td>
<td>Stopped by aluminium</td>
</tr>
<tr>
<td>Gamma $\gamma$</td>
<td>Electromagnetic waves</td>
<td>Great distances</td>
<td>Weak</td>
<td>Stopped by thick lead</td>
</tr>
</tbody>
</table>

**Gamma rays** can only effectively be stopped by a very thick piece of lead (Figure 18.13). Gamma rays only ionise air very weakly and travel great distances through air.

![Figure 18.13](image.png) The penetrating power of alpha, beta and gamma radiation.

**Test yourself**

20 Which one of the following, A, B or C, is a property of beta radiation?
   A It is the most strongly ionising radiation.
   B It will travel through several metres of air.
   C It can be easily stopped by paper.

21 a) Which type of radiation has a range of a few centimetres in air?
   b) Which types of radiation are stopped by thin metal sheets?

22 Explain why a teacher uses long tongs to handle a source of radiation.

**Radioactive decay**

The atoms of some radioactive materials decay by emitting alpha, beta or gamma radiations from their nuclei. However, it is not possible to predict when the nucleus of one particular atom will decay. It could be the next second or sometime next week or not for a million years. Radioactive decay is a random process.

**Random process**

The radioactive decay of an atom is rather like rolling dice or tossing a coin. You cannot say with certainty that the next time you toss a coin it will fall heads up. However, if you throw a lot of coins you can start to predict how many of them will fall heads up. You can use this idea to help you understand how radioactive decay happens. Imagine you start off with a thousand coins. When any coin falls heads up then it has ‘decayed’ and you take it out of the game. Table 18.2 shows the likely result (on average). Every time you throw a lot of coins, about half of them will turn up heads.

![Figure 18.14](image.png) On average how many sixes will turn up when you roll 100 dice? Why can you not be certain what will happen on each occasion?
Radioactive decay

Radioactive materials decay in a similar way. If we start off with a million atoms then after a period of time (for example 1 hour), half of them will have decayed. In the next hour we find that half of the remaining atoms have decayed, leaving us with a quarter of the original number. The period of time taken for half the number of nuclei to decay in a radioactive sample is called the half-life and it is given the symbol \( t_{1/2} \).

It is important to understand that we have chosen a half-life here of 1 hour to explain the idea. Different radioisotopes have different half-lives.

Count rate and activity

The activity of a source is equal to the number of particles emitted per second. We can express this as counts per second, but in honour of Henry Becquerel, this unit is called the becquerel (Bq).

The count rate is the term we use when a GM tube is measuring the radiation emitted from a radioactive source. This is different from the activity of a source, because not all the radiation emitted from the source goes into the GM tube. Count rates may be in counts per second or sometimes counts per minute.

Measurement of half-life

If you look at Table 18.5, you can see that the number of nuclei that decayed in the first hour was 500 000, then in the next hour 250 000 and in the third hour 125 000. So as time passes not only does the number of nuclei left get smaller but so does the rate at which the nuclei decay. So by measuring the activity of a radioactive sample we can determine its half-life.

Figure 18.15 shows the result of a laboratory experiment to determine the half-life of a radioactive material. You can see that the count rate detected halves every half-life. At the start of the experiment the count rate was 40 per second, after 50 seconds (one half-life) it has reduced to 20 per second, and after a further 50 seconds the count rate has halved again to 10 per second.
Test yourself

23 Use a word from the box to complete the sentence.

- count rate
- half-life
- reaction

The _______ is the number of alpha, beta or gamma emissions detected from a radioactive source in one second.

24 The graphs in Figure 18.16 show the decay of three different radioactive isotopes. Which isotope has:

   a) the longest half-life?
   b) the shortest half-life?

25 Explain what the word random means.

26 A radioactive material has a half-life of 15 minutes. What does this mean? How much of the original material will be left after 60 minutes?

27 The results in the table opposite for the count rate of a radioactive source were recorded every minute. Plot a graph of the count rate (y-axis) against time (x-axis) and use the graph to work out the half-life of the source.

28 A radioactive isotope has a half-life of 8 hours. At 12 noon on 2 March a GM tube measures a count rate of 2400 per second.

   a) What will be the count rate at 4.00 am on 3 March?
   b) At what time will a count rate of approximately 75 per second be measured?

---

Table 18.6 Half-lives of some radioisotopes.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Half-life</th>
</tr>
</thead>
<tbody>
<tr>
<td>Potassium-40</td>
<td>1.3 billion years</td>
</tr>
<tr>
<td>Carbon-14</td>
<td>5700 years</td>
</tr>
<tr>
<td>Caesium-137</td>
<td>30 years</td>
</tr>
<tr>
<td>Iodine-131</td>
<td>8 days</td>
</tr>
<tr>
<td>Lawrencium-260</td>
<td>3 minutes</td>
</tr>
<tr>
<td>Nobelium-252</td>
<td>2.3 seconds</td>
</tr>
</tbody>
</table>

---

Table 18.7

<table>
<thead>
<tr>
<th>Count rate in Bq</th>
<th>Time in minutes</th>
</tr>
</thead>
<tbody>
<tr>
<td>1000</td>
<td>0</td>
</tr>
<tr>
<td>590</td>
<td>1</td>
</tr>
<tr>
<td>340</td>
<td>2</td>
</tr>
<tr>
<td>200</td>
<td>3</td>
</tr>
<tr>
<td>120</td>
<td>4</td>
</tr>
<tr>
<td>70</td>
<td>5</td>
</tr>
</tbody>
</table>
Radiation damage

If radiation gets into our body, damage can occur to cells and tissues. The ions which are produced by the radiation produce chemicals which destroy cells they come into contact with. Alpha particles cause the most damage if they get inside the body, as they are strongly ionising; this could happen if we breathed in a radioactive gas. An alpha source in school is less dangerous as the radiation only travels short distances, so does not enter our body.

Although gamma rays are less ionising than alpha particles, a gamma ray source in a laboratory is a hazard, as the rays are so penetrating. A gamma ray can pass into our body from a source several metres away from us.

Table 18.8 gives some examples of when we are exposed to radiation and the risk.

Low doses

Low doses of radiation, are unlikely to cause us harm. However, it is possible that any exposure to radiation will increase our chances of cancer.

Moderate doses

Moderate doses of radiation are unlikely to kill someone, but the person will be very unwell. Damage will be done to cells in the body, but not enough to be fatal. The body will be able to replace dead cells and the person is likely to recover completely. However, studies of the survivors from the Hiroshima and Nagasaki bombs in 1945, and of survivors from the Chernobyl reactor disaster of 1986, show that there is an increased chance of dying from cancer some years after the dose of radiation.

High doses

High doses of radiation are likely to be fatal. A high dose damages the gut and bone marrow so much that the body cannot work normally. About 30 people died of acute radiation syndrome in the Chernobyl disaster.

Table 18.8 Radiation doses and their effects.

<table>
<thead>
<tr>
<th>Source of dose</th>
<th>Effect of dose</th>
</tr>
</thead>
<tbody>
<tr>
<td>Airport security scan; Eating a banana</td>
<td>Low risk</td>
</tr>
<tr>
<td>Dental X-ray</td>
<td>Low risk</td>
</tr>
<tr>
<td>Transatlantic flight</td>
<td>Low risk</td>
</tr>
<tr>
<td>Chest X-ray</td>
<td>Low risk</td>
</tr>
<tr>
<td>Release limit from a nuclear plant per person per year</td>
<td>Low risk</td>
</tr>
<tr>
<td>Yearly dose from food</td>
<td>Low risk</td>
</tr>
<tr>
<td>Mammogram</td>
<td>Low risk</td>
</tr>
<tr>
<td>Average computer tomography (CT) scan</td>
<td>Low risk</td>
</tr>
<tr>
<td>Maximum yearly dose permitted for radiation workers</td>
<td>Medium risk</td>
</tr>
<tr>
<td>Lowest annual dose where increased risk of cancer is evident</td>
<td>Medium risk</td>
</tr>
<tr>
<td>Highly targeted dose used in radiotherapy (single dose)</td>
<td>High risk, but balanced by a likely cure of cancer</td>
</tr>
<tr>
<td>Extremely severe dose, received in a nuclear accident</td>
<td>Death probable within 6 weeks</td>
</tr>
<tr>
<td>Maximum radiation dose per day found at the Fukushima plant in 2011</td>
<td>Fatal dose; death within 2 weeks</td>
</tr>
<tr>
<td>10 minutes exposure to the Chernobyl reaction meltdown in 1986</td>
<td>Death within hours</td>
</tr>
</tbody>
</table>
Name four nuclear radiations which are dangerous to us.

The table below shows the results of research into the number of deaths caused by various types of radiation.

<table>
<thead>
<tr>
<th>Source of radiation</th>
<th>Type of radiation</th>
<th>Number of people studied</th>
<th>Extra number of cancer deaths caused by radiation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Uranium miners</td>
<td>Alpha</td>
<td>3400</td>
<td>60</td>
</tr>
<tr>
<td>Radium luminisers</td>
<td>Alpha</td>
<td>800</td>
<td>50</td>
</tr>
<tr>
<td>Medical treatment</td>
<td>Alpha</td>
<td>4500</td>
<td>60</td>
</tr>
<tr>
<td>Hiroshima bomb</td>
<td>Gamma rays and neutrons</td>
<td>15,000</td>
<td>100</td>
</tr>
<tr>
<td>Nagasaki bomb</td>
<td>Gamma rays</td>
<td>7000</td>
<td>20</td>
</tr>
</tbody>
</table>

a) Discuss whether the table supports the suggestion that alpha particles are more dangerous than gamma rays.

b) What conclusion can you draw about the relative dangers of neutrons?

c) A student comments that the table does not provide a fair test for comparison of the radiations. Comment on this.

Complete this task to show that you understand the hazards of radiation.

Explain how radiation can damage our bodies. Explain how measuring radiation doses can help reduce risk to people.

○ Irradiation and contamination

When a teacher brings a radioactive source into a laboratory to demonstrate to her pupils, she irradiates the surroundings (with a very small safe dose) but does not cause any contamination. When the source is near a GM tube, the tube is irradiated. This means that radiation from the source is entering the tube. However, as soon as the source is removed and put away, the GM tube will not be radioactive, because there are no radioactive nuclei inside it to emit radiation.

Radioactive isotopes cause contamination when they get into places where we do not want them. For example, iodine-131 and caesium-137 were emitted in the Fukushima and Chernobyl disasters. Iodine-131 has a half-life of 8 days, which means that it is highly active for a short time. Iodine disperses in the atmosphere and gets into the food chain. This poses a very high risk, but for a short period. Caesium-137, however, has a half-life of 30 years. Consequently, the ground and water in the region close to the two nuclear reactors will be contaminated for many years to come.

Both irradiation and contamination are potentially hazardous to us. However, when a patient is exposed to a dose of radiation, for example, in hospital, the dose will be carefully calculated and controlled. The problem with contamination is that a person is exposed to radiation in a way which is unknown, and it is possible that a large and very dangerous dose of radiation is consumed by mistake.

Injecting a radioactive substance into a person to diagnose a medical problem always involves some risk to a person's health. Use an answer from the box to complete the sentence.

greater than  the same as  less than

A radioactive substance may be used to diagnose a medical problem if the potential benefit of the diagnosis is ________ than the risk to the person's health.
Chapter review questions

1. Lithium atoms have three protons and four neutrons.
   a) What is the atomic number of lithium?
   b) What is the mass number of lithium?
   c) How many electrons are there in a lithium atom?

2. There are two stable isotopes of carbon: carbon-12 and carbon-13.
   a) Explain what is meant by the words
      i) stable
      ii) isotope.
   b) The atomic number of carbon is 6. How many protons and neutrons are there in each of the isotopes mentioned above?

3. How do the atoms of one particular element differ in atomic structure from the atoms of all other elements?

4. a) Write nuclear equations for the alpha decay of:
     i) \(^{241}_{94}\)Pu to uranium (U)
     ii) \(^{229}_{90}\)Th to radium (Ra)
     iii) \(^{213}_{84}\)Po to lead (Pb).
   b) Write nuclear equations for the beta decay of:
     i) \(^{237}_{92}\)U to neptunium (Np)
     ii) \(^{59}_{26}\)Fe to cobalt (Co)
     iii) \(^{32}_{14}\)Si to phosphorus (P).

5. Why did the work done by Geiger and Marsden convince scientists that the ‘plum pudding’ model of the atom needed to be replaced?

6. In the alpha scattering experiment about 1 in 10,000 alpha particles bounced back from the gold foil. Explain how you think the number of alpha particles bouncing back will change when:
   a) thicker gold foil is used
   b) aluminium foil of the same thickness is used.

   a) What is an alpha particle?
   b) Explain why bismuth-213 would be highly dangerous if put inside the body.
8. Zak’s teacher carried out an experiment to measure the half-life of protactinium-234. His results are shown in the table.

Table 18.10

<table>
<thead>
<tr>
<th>Time in seconds</th>
<th>Count rate in counts per second</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>66</td>
</tr>
<tr>
<td>40</td>
<td>44</td>
</tr>
<tr>
<td>80</td>
<td>30</td>
</tr>
<tr>
<td>120</td>
<td>20</td>
</tr>
<tr>
<td>160</td>
<td>13</td>
</tr>
<tr>
<td>200</td>
<td>9</td>
</tr>
<tr>
<td>240</td>
<td>6</td>
</tr>
<tr>
<td>280</td>
<td>4</td>
</tr>
</tbody>
</table>

a) In the experiment, what was:
   i) the independent variable
   ii) the dependent variable?

b) What instrument was used to measure the count rate?

9. The table gives information about some radioactive sources that emit ionising radiation.

Table 18.11

<table>
<thead>
<tr>
<th>Source</th>
<th>Radiation emitted</th>
<th>Half-life</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bismuth-213</td>
<td>Alpha</td>
<td>45 minutes</td>
</tr>
<tr>
<td>Cobalt-60</td>
<td>Gamma</td>
<td>5 years</td>
</tr>
<tr>
<td>Uranium-233</td>
<td>Alpha</td>
<td>150,000 years</td>
</tr>
<tr>
<td>Radon-226</td>
<td>Beta</td>
<td>6 minutes</td>
</tr>
<tr>
<td>Technetium-99</td>
<td>Gamma</td>
<td>6 hours</td>
</tr>
</tbody>
</table>

a) What is ‘half-life’?

b) i) Explain what is meant by the term ‘ionising radiation’.
   ii) Which of the radiations shown in the table is the most ionising?
**Practice questions**

1 The diagram represents an atom of beryllium-9.

![Figure 18.17](image)

**a)** i) Copy and complete the following table of information for an atom of beryllium-9.  [3 marks]

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Number of electrons</td>
<td></td>
</tr>
<tr>
<td>Number of protons</td>
<td></td>
</tr>
<tr>
<td>Number of neutrons</td>
<td></td>
</tr>
</tbody>
</table>

**ii)** What is the atomic number of a beryllium-9 atom?

**Choose the correct answer from the box.**

4 5 9 13  [2 marks]

Give the reason for your answer.

**b)** Beryllium-10 is a radioactive isotope of beryllium.

i) **Choose the correct answer from the box to complete the sentence.**  [1 mark]

- electron
- neutron
- proton

The nucleus of a beryllium-10 atom has one more ________ than the nucleus of a beryllium-9 atom.

**ii)** Beryllium-10 decays by emitting beta particles.

Which statement, A, B or C, describes a beta particle?  [1 mark]

- A the same as a helium nucleus
- B an electromagnetic wave
- C an electron from the nucleus

**c)** The graph shows how the count rate from a sample of beryllium-10 changes with time.

![Figure 18.18](image)

i) How many millions of years does it take for the count rate to fall from 400 Bq to 40 Bq?  [1 mark]

ii) What is the half-life of beryllium-10?  [1 mark]

2 **a)** The statements A, B and C give three properties of nuclear radiation.

- A It will pass through cardboard but not thin aluminium.
- B It is weakly ionising.
- C It can travel only a few centimetres through the air.

i) Which **one** of the statements, A, B or C, gives a property of alpha radiation?  [1 mark]

**ii)** Which **one** of the statements, A, B or C, gives a property of gamma radiation?  [1 mark]

**b)** Fresh strawberries grown abroad are sometimes irradiated before being sent to the UK.

The irradiation process kills bacteria on the strawberries.

i) **Which one** of the statements, X, Y or Z, is correct?  [1 mark]

- X The irradiated strawberries do not become radioactive.
- Y Particles containing radioactive atoms settle on the strawberries.
- Z The strawberries cannot be eaten for a few days after irradiation.

**ii)** Suggest one reason why the farmers growing the strawberries want the strawberries to be irradiated.  [1 mark]

3 **a)** Phosphorus is an element with an atomic number of 15. Its most common isotope is phosphorus-31, $^{31}\text{P}$. Another isotope, phosphorus-32, is radioactive.

i) State the number of protons, neutrons and electrons in phosphorus-31.  [2 marks]

ii) Why does phosphorus-31 have a different mass number from phosphorus-32?  [1 mark]

**iii)** Atoms of phosphorus-32 change into atoms of sulfur by beta decay. Copy and complete the equation to show the atomic and mass numbers of this isotope of sulfur.

$^{32}\text{P} \rightarrow \text{S} + \text{beta} \quad \text{[2 marks]}$

**b)** i) Name a suitable detector that could be used to show that phosphorus-32 gives out radiation.  [1 mark]

**ii)** Name a disease that can be caused by too much exposure to a radioactive substance such as phosphorus-32.  [1 mark]
4 A radioactive source emits alpha (α), beta (β) and gamma (γ) radiation.

a) Which two types of radiation will pass through a sheet of card? [1 mark]

b) Which type of radiation has the greatest range in air? [1 mark]

5 The table gives information about some of the radioactive substances released into the air by the explosion at the Fukushima nuclear plant in 2011.

<table>
<thead>
<tr>
<th>Radioactive substance</th>
<th>Half-life</th>
<th>Type of radiation emitted</th>
</tr>
</thead>
<tbody>
<tr>
<td>Iodine-131</td>
<td>8 days</td>
<td>Beta and gamma</td>
</tr>
<tr>
<td>Caesium-134</td>
<td>2 years</td>
<td>Beta</td>
</tr>
<tr>
<td>Caesium-137</td>
<td>30 years</td>
<td>Beta</td>
</tr>
</tbody>
</table>

a) How is the structure of a caesium-134 atom different from the structure of a caesium-137 atom? [1 mark]

b) What are beta and gamma radiations? [2 marks]

c) A sample of soil is contaminated with some iodine-131. Its activity is 10000 Bq. Calculate how long it will take for the activity to drop to 2500 Bq. [2 marks]

d) Which of the three isotopes will be the most dangerous 50 years after the accident? Explain your answer. [2 marks]

6 The table gives information about five radioactive isotopes.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Radiation emitted</th>
<th>Half-life</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>Alpha</td>
<td>4 minutes</td>
</tr>
<tr>
<td>B</td>
<td>Gamma</td>
<td>5 years</td>
</tr>
<tr>
<td>C</td>
<td>Beta</td>
<td>12 years</td>
</tr>
<tr>
<td>D</td>
<td>Beta</td>
<td>28 years</td>
</tr>
<tr>
<td>E</td>
<td>Gamma</td>
<td>6 hours</td>
</tr>
</tbody>
</table>

a) What is a ‘beta particle’? [1 mark]

b) What does the term ‘half-life’ mean? [1 mark]

c) Radioactive waste needs to be stored. One suggestion is to seal it in steel drums and bury these in deep underground caverns. Suggest why people may be worried about having such a storage site close to where they live. [3 marks]

7 Some patients who suffer from cancer are given an injection of boron, which is absorbed by cancer cells. The cancerous tissue is then irradiated with neutrons. After this, the reaction that occurs is:

\[ _{7}^{12}B \rightarrow _{2}^{3}Li + _{2}^{3}He \]

a) Copy the equation and write numbers in place of x and y. [3 marks]

b) Why do the lithium and helium nuclei repel each other? [1 mark]

c) Why is this process dangerous for healthy patients? [1 mark]

8 In the early 20th century scientists thought that atoms were made up of electrons embedded inside a ball of positive charge. They called this the ‘plum pudding’ model of the atom.

Geiger and Marsden fired a beam of alpha particles at a thin gold foil. Explain how the results of their experiment led to a new model of the atom. Illustrate your answer with suitable diagrams. [6 marks]
Working scientifically: Risk and perception of risk

Before starting an investigation you complete a risk assessment. You identify the hazards, the risks and the controls needed to reduce the identified risks. For example, if you use boiling water there is always a risk you may scald yourself.

1 Look at the information on radiation and on radiation damage earlier in this chapter. What controls should be used to reduce the risk to the teacher and students of using a radioactive source?

Perception of risk
The perception of risk is often different from the actual risk. This happens for different reasons.

Voluntary versus imposed risk
When people are doing an activity by choice, they may perceive it to have a smaller risk than if they have no choice in doing something. For example, some people voluntarily throw themselves from high bridges and cranes. They call it bungee jumping!

Familiar versus unfamiliar risk
People may be happy to use electrical appliances but worry about the risk of a nuclear reactor accident. In fact, many more people have been killed or injured by electrical accidents in the home than by accidents at nuclear reactors.

Visible versus invisible hazard
Every day we accept the risks involved in crossing the road. The hazards are real but our perception of the risk is likely to be reduced because generally we can see the hazards. Nuclear radiation, however, is invisible. We can’t see it, we can’t feel it, we can’t smell it. The perceived risk is usually high.

2 Which of the following involves a voluntary risk?
   a) Flying to a holiday destination.
   b) Riding a bicycle to school.
   c) Exposure to background radiation.
   d) Learning to drive a car.

Electrical safety in the home
A recent survey found that many people perceive the risk of injury from using electrical appliances to be much lower than the accident figures suggest. On average, fires caused by the misuse of electrical appliances kill one person each week and seriously injure about 2500 people each year. In addition, thousands of people are injured each year (with about 30 deaths) by electric shocks. Perhaps it’s because electrical appliances are so familiar to us that we simply forget or underestimate the risks.
**Appendix: The periodic table**

<table>
<thead>
<tr>
<th>Key</th>
<th>Relative atomic mass</th>
<th>Atomic symbol</th>
<th>Atomic (proton) number</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>H</td>
<td>hydrogen</td>
<td>1</td>
<td></td>
</tr>
<tr>
<td>7</td>
<td>Li</td>
<td>lithium</td>
<td>3</td>
<td></td>
</tr>
<tr>
<td>9</td>
<td>Be</td>
<td>beryllium</td>
<td>4</td>
<td></td>
</tr>
<tr>
<td>11</td>
<td>Na</td>
<td>sodium</td>
<td>11</td>
<td></td>
</tr>
<tr>
<td>12</td>
<td>Mg</td>
<td>magnesium</td>
<td>12</td>
<td></td>
</tr>
<tr>
<td>13</td>
<td>Al</td>
<td>aluminium</td>
<td>13</td>
<td></td>
</tr>
<tr>
<td>14</td>
<td>Si</td>
<td>silicon</td>
<td>14</td>
<td></td>
</tr>
<tr>
<td>15</td>
<td>P</td>
<td>phosphorus</td>
<td>15</td>
<td></td>
</tr>
<tr>
<td>16</td>
<td>S</td>
<td>sulphur</td>
<td>16</td>
<td></td>
</tr>
<tr>
<td>17</td>
<td>Cl</td>
<td>chlorine</td>
<td>17</td>
<td></td>
</tr>
<tr>
<td>18</td>
<td>Ar</td>
<td>argon</td>
<td>18</td>
<td></td>
</tr>
<tr>
<td>19</td>
<td>K</td>
<td>potassium</td>
<td>19</td>
<td></td>
</tr>
<tr>
<td>20</td>
<td>Ca</td>
<td>calcium</td>
<td>20</td>
<td></td>
</tr>
<tr>
<td>21</td>
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<td>21</td>
<td></td>
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<tr>
<td>22</td>
<td>Ti</td>
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<td>23</td>
<td>V</td>
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<td>Cr</td>
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<td>25</td>
<td>Mn</td>
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<td>26</td>
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<td>27</td>
<td>Co</td>
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<td>28</td>
<td>Ni</td>
<td>nickel</td>
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<tr>
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<td>30</td>
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<td>Ga</td>
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<td>31</td>
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<tr>
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<td>35</td>
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<td>36</td>
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<td>krypton</td>
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<td>Rb</td>
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<td>Ru</td>
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<td>Rh</td>
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<td>Pd</td>
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<td>In</td>
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<td></td>
</tr>
<tr>
<td>50</td>
<td>Sn</td>
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<td>Pb</td>
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<td>Bi</td>
<td>bismuth</td>
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<td>radon</td>
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<td>Sg</td>
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* The Lanthanides (atomic numbers 58 – 71) and the Actinides (atomic numbers 90 – 103) have been omitted.

Relative atomic masses for **Cu** and **Cl** have not been rounded to the nearest whole number.
Glossary

**Accurate** A measurement or calculated value that is close to the true value.

**Acid** Solution with a pH less than 7; produces H+ ions in water.

**Acid rain** Precipitation that is acidic as a result of air pollution.

**Activation energy** The minimum energy particles must have to react.

**Active site** The region of an enzyme that binds to its substrate.

**Active transport** The net movement of particles from an area of low concentration to an area of higher concentration using energy.

**Aerobic** In the presence of oxygen.

**Aerobic respiration** Respiration using oxygen.

**Alkali** Solution with a pH more than 7; produces OH− ions in water.

**Alkali metals** The elements in Group 1 of the periodic table (including lithium, sodium and potassium).

**Alloy** A mixture of a metal with small amounts of other elements, usually other metals.

**Alpha particle** A particle formed from two protons and two neutrons.

**Alternative energy resources** Resources other than fossil fuels.

**Antimicrobial secretions** Chemicals that destroy pathogens.

**Antibiotic** A result that does not fit the expected pattern.

**Antibiotic-resistant bacterium** A bacterium that cannot be killed by antibiotics.

**Anode** An electrode where oxidation takes place (oxidation is the loss of electrons) – in electrolysis, it is the positive electrode.

**Anomalous** A result that doesn’t fit the expected pattern.

**Anomalous result** A result that does not fit the expected pattern.

**Antibiotic** A group of medicines, first discovered by Sir Alexander Fleming, that kill bacteria and fungi but not viruses.

**Antibiotic-resistant bacterium** A bacterium that cannot be killed by antibiotics.

**Antibody** A protein produced by lymphocytes that recognises pathogens and helps to clump them together.

**Antigen** A protein on the surface of a pathogen that antibodies can recognise as foreign.

**Antimicrobial secretions** Chemicals that destroy pathogens.

**Antitoxin** A protein produced by your body to neutralise harmful toxins produced by pathogens.

**Atom** The smallest part of an element that can exist. A particle with no electric charge made up of a nucleus containing protons and neutrons surrounded by electrons in energy levels.

**Atomic number** Number of protons in an atom.

**Atrium (plural atria)** An upper chamber of the heart surrounded by a thin wall of muscle.

**Avogadro constant** The number of atoms, molecules or ions in one mole of a given substance (the value of the Avogadro constant is 6.02 × 10^23).

**Axon** The extension of a nerve cell along which electrical impulses travel.

**Battery** Two or more chemical cells connected together.

**Benign** A non-cancerous tumour that does not spread.

**Beta particle** A fast moving electron.

**Bile** A green-coloured liquid produced by your liver, stored by your gall bladder and released into your small intestine to help break down fats.

**Biofuel** Fuel produced from biological material. Biofuels are provided by trees such as willow that can be grown specifically as energy resources.

**Blood plasma** The straw-coloured liquid that carries our blood cells and dissolved molecules.

**Braking distance** The distance a car travels while the car is stopped by the brakes.

**Burette** A glass tube with a tap and scale for measuring liquids to the nearest 0.1 cm³.

**Calibrate** Mark a scale onto a measuring instrument so that you can give a value to a measured quantity.

**Capillaries** Tiny blood vessels found between arteries and veins that carry blood into tissues and organs.

**Carbohydrates** Biological molecules containing carbon, hydrogen and oxygen.

**Carbon neutral** Fuels and processes whose use results in zero net release of greenhouse gases to the atmosphere.

**Carcinogen** A cancer-causing substance.

**Catalyst** Substance that speeds up a chemical reaction but is not used up.

**Categorical variable** A variable with values that are given a name or label.

**Cathode** An electrode where reduction takes place (reduction is the gain of electrons) – in electrolysis it is the negative electrode.

**Causation** The act of making something happen.

**Cell** Two electrodes in an electrolyte used to generate electricity.

**Cell cycle** A series of steps in which the genetic material doubles and the cell divides into two identical cells.

**Chemical bond** The attraction between atoms that holds them together to form molecules.

**Cholesterol** An important biological molecule for cell membranes but leads to atherosclerosis if found in high levels in the blood.

**Chromosome** Structure containing DNA, found in the nucleus of eukaryotic cells.

**Cilia** Tiny hair-like projections from ciliated cells that waft mucus out of the gas exchange system.

**Clone** An organism produced asexually that has identical genes to its parent.

**Cloning** The asexual reproduction of an organism to produce genetically identical offspring.

**Combustion** Burning.

**Common ancestor** An organism from which others have evolved.

**Communicable** A disease that can be transmitted from one organism to another.

**Compound** Substance made from different elements chemically bonded together.

**Computer modelling** Using computer software to theoretically examine or test.
Concentrated A solution in which there is a lot of solute dissolved.  
Concentration gradient A measurement of how a concentration of a substance changes from one place to another.  
Condenser The part of the apparatus that causes the solvent to condense to a liquid.  
Conservation of mass In a reaction, the total mass of the reactants must equal the total mass of the products.  
Conservation Protecting an ecosystem or species of organism from reduced numbers and often extinction.  
Continuous (data) Data that come in a range and not in groups.  
Continuous variable A variable with numerical values obtained by either measuring or counting.  
Control variable This is the variable that can affect the outcome of an investigation and therefore must be kept constant or monitored. If all control variables are kept constant then the experiment is a fair test.  
Coronary arteries Arteries that supply the heart muscle with oxygenated blood.  
Correlation When an action and outcome are linked but the action does not necessarily cause the outcome.  
Covalent bond Two shared electrons joining atoms together.  
Daughter cells The cells produced during mitosis.  
Defecation Removing solid waste from the body.  
Deficiency A lack or shortage.  
Delocalised Free to move around.  
Denatured A permanent change to an enzyme as a result of extremes of pH and temperature, which stop it working.  
Deoxygenated Without oxygen.  
Dependent variable This is the variable that is measured or recorded for each change of the independent variable.  
Diabetes A non-communicable disease that reduces control of blood glucose concentrations.  
Diatomic molecule A molecule containing two atoms.  
Differentiate To specialise, or adapt for a particular function.  
Diffusion The net movement of particles from an area of high concentration to an area of lower concentration.  
Dilute A solution in which there is a small amount of solute dissolved.  
Diploid Describes a cell or nucleus of a cell that has a paired set of chromosomes.  
Directly proportional When two quantities are directly proportional, doubling one quantity will cause the other quantity to double; when a graph is plotted, the graph line will be straight and pass through the origin (0, 0).  
Discharge Gain or lose electrons to become electrically neutral.  
Displacement A distance travelled in a defined direction.  
Displacement reaction Reaction where a more reactive element takes the place of a less reactive element in a compound.  
Dissipate To scatter in all directions or to use wastefully; the energy has spread out and heats up the surroundings.  
DNA (deoxyribonucleic acid) The genetic information found in all living organisms.  
Double blind trial A medical experiment in which the patients and doctors do not know who has been given the drug and who has been given the placebo.  
Efficacy How effective a drug is.  
Electric current A flow of electrical charge; the size of the electric current is the rate at which electrical charge flows round the circuit.  
Electrolysis Decomposition of ionic compounds using electricity.  
Electrolyte A liquid that conducts electricity.  
Electron A negatively charged particle that orbits the nucleus of an atom.  
Electron microscope A microscope that uses electron beams in place of light to give higher magnification.  
Electronic structure The arrangement of electrons in the shells (energy levels) of an atom.  
Element A substance containing only one type of atom; a substance that cannot be broken down into simpler substances by chemical methods.  
Endothermic reaction Reaction where thermal energy is transferred from the surroundings to the chemicals and so the temperature decreases.  
Energy level (shell) The region an electron occupies surrounding the nucleus inside an atom.  
Environmental variation Differences in organisms as a result of the environment in which they live.  
Enzymes Molecules that act as catalysts in biological systems.  
Epidermis The outermost layer of cells of a plant.  
Ethene A plant hormone that ripens fruit.  
Ethical issue An idea some people disagree with for religious or moral reasons.  
Eukaryote An organism that is made of eukaryotic cells (those that contain a nucleus).  
Eukaryotic cells Cells that contain a nucleus.  
Evolution The theory first proposed by Charles Darwin that the different species found today formed as a result of the accumulation of small advantages that were passed through generations.  
Excess When the amount of a reactant is greater than the amount that can react.  
Excretion The removal of substances produced by chemical reactions inside cells from cells or organisms.  
Exothermic reaction Reaction where thermal energy is transferred from the chemicals to the surroundings and so the temperature increases.  
Extension The difference between the stretched and unstretched lengths of a spring.  
Extinct When no members of a species remain alive.  
Extrapolation (or prediction) Assuming that an existing trend or pattern continues to apply in an unknown situation.  
Fair test Only the independent variable affects the dependent variable.  
Fermentation The chemical breakdown of glucose into ethanol and carbon dioxide by respiring microorganisms such as yeast.  
Filterate Liquid that comes through the filter paper during filtration.  
Fluid A liquid or a gas; a fluid flows and can change shape to fill any container.  
Force A push or a pull.  
Fraction A mixture of molecules with similar boiling points.  
Fractional distillation A method used to separate miscible liquids with different boiling points.  
Frequency (Hz) The number of waves produced each second or the number of waves passing a point each second. The unit of frequency is hertz; 1 hertz means there is 1 cycle per second.  
Fullerenes Family of carbon molecules each with carbon atoms linked in rings to form a hollow sphere or tube.  
Gametes Sex cells, e.g. sperm, ova and pollen.  
Gamma ray An electromagnetic wave.  
Geiger-Müller (GM) tube A device which detects ionising radiation; an electronic counter can record the number of particles entering the tube.
**Gene** A section of a chromosome made from DNA that carries the code to make a protein.

**Genetic variation** Inherited differences in organisms.

**Giant lattice** A regular structure containing a massive number of particles that continue in all directions throughout the structure.

**Global warming** An increase in the temperature at the Earth’s surface.

**Glycogen** An insoluble store of glucose in the liver.

**Greenhouse gas** A gas that absorbs long wavelength infrared radiation given off by the Earth but does not absorb the Sun’s radiation.

**Group** The name given to each column in the periodic table.

**Haemoglobin** The molecule in red blood cells that can temporarily bind with oxygen to carry it around your body.

**Half-life** The average time taken for the number of nuclei in a radioactive isotope to halve; in one half-life the activity or count rate of a radioactive sample also halves.

**Halides** Compounds made from Group 7 elements.

**Halogen** The elements in Group 7 of the periodic table (including fluorine, chlorine, bromine and iodine).

**Haploid** Describes a cell or nucleus of a gamete that has an unpaired set of chromosomes (i.e. only half the normal number).

**Hazard** Something that could cause harm.

**Heart bypass** A medical procedure in which a section of less important artery is moved to allow blood to flow around a blockage in a more important artery.

**Hormone** A chemical (produced in a gland in mammals) that moves around an organism to change the function of target cells, tissues or organs.

**Humid** Describes an atmosphere with high levels of water vapour.

**Hydrocarbon** A compound containing hydrogen and carbon only.

**Hypothesis** An idea that explains how or why something happens.

**Immiscible** Liquids that do not mix together and separate into layers.

**Independent variable** This is the variable that is changed in an experiment or selected by the investigator.

**Inert electrodes** Electrodes that allow electrolysis to take place but do not react themselves.

**Infectious** Describes a pathogen that can easily be transmitted, or an infected person who can pass on the disease.

**Insecticide** A chemical that kills insects.

**Insoluble** Cannot dissolve.

**Insulin** A hormone produced in your pancreas that lowers blood glucose by converting it to glycogen and storing it in the liver.

**Interval** The difference between one value in a set of data and the next.

**Inversely proportional** When two quantities are inversely proportional, doubling one quantity will cause the other quantity to halve.

**In vitro fertilisation (IVF)** A medical procedure in which ova are fertilised outside of a woman, then placed into her uterus to develop into a baby.

**Ion** An electrically charged particle containing different numbers of protons and electrons.

**Ionic bonding** The electrostatic attraction between positive and negative ions.

**Ionic equation** Balanced equation for reaction that omits any spectator ions.

**Ionising radiation** UV rays, X-rays and gamma rays that can cause mutations to DNA.

**Isotopes** Different forms of a particular element; isotopes have the same number of protons but different numbers of neutrons.

**Kingdom** The largest group of classifying organisms, e.g. the animal kingdom.

**Limiting factor** Anything that reduces or stops the rate of a reaction.

**Limiting reactant** The reactant in a reaction that determines the amount of products formed. Any other reagents are in excess and will not all react.

**Lipids** Fats or oils, which are insoluble in water.

**Lock and key hypothesis** A model that explains the action of enzymes.

**Lymphocyte** A type of white blood cell that produces antibodies to help clump pathogens together to make them easier to destroy.

**Malaria** A communicable disease, caused by a protist transmitted in mosquitos, which attacks red blood cells.

**Malignant** A cancerous tumour that can spread to other parts of the body.

**Malleable** Can be hammered into shape.

**Mass number** Number of protons plus the number of neutrons in an atom.

**Meniscus** The curve at the surface of a liquid in a container.

**Metabolism** The sum of all the chemical reactions that happen in an organism.

**Metallic bonding** The attraction between the nucleus of metal atoms and delocalised electrons.

**Mineral ions** Substances that are essential for healthy plant growth, e.g. nitrates and magnesium.

**Miscible** Liquids that mix together.

**Mitochondrion** A small cell organelle in which respiration occurs, found in the cytoplasm of eukaryotic cells.

**Mitosis** Cell replication that produces two identical copies of a diploid cell.

**Mixture** More than one substance that are not chemically joined together.

**Mole** Measurement of the amount of a substance.

**Molecule** Particle made from atoms joined together by covalent bonds.

**Monomer** The building block (molecule) of a polymer.

**Mucus** A sticky substance that traps pathogens.

**Mutation** A permanent change to DNA, which may be advantageous, disadvantageous or have no effect.

**Myelin sheath** The insulating cover along an axon, which speeds up the electrical impulse.

**Nanoscience** The study of nanoparticles.

**Net** Overall.

**Neutralisation** A reaction that uses up some or all of the H+ ions from an acid.

**Neutron** A neutral particle found in the nucleus of an atom.

**Noble gases** The elements in Group 0 of the periodic table (including helium, neon and argon).
Non-communicable disease A disease that is not passed from person to person.
Non-ohmic The current flowing through a non-ohmic resistor is not proportional to the potential difference across it; the resistance changes as the current flowing through it changes.
Non-renewable energy resources Energy resources which will run out and cannot be replenished.
Nucleus Central part of an atom containing protons and neutrons.
Ohmic The current flowing through an ohmic conductor is proportional to the potential difference across it; if the p.d. doubles, the current doubles but he resistance stays the same.
Ore A rock from which a metal can be extracted for profit.
Organelle A part of a cell with a specific function.
Osmosis The net diffusion of water from an area of high concentration of water to an area of lower concentration of water across a partially permeable membrane.
Ova (singular ovum) Eggs.
Oxidation A reaction that uses oxygen.
Oxygen debt The amount of extra oxygen the body needs after exercise to break down the lactic acid.
Oxygenated Rich in oxygen.
Oxymoglobin The name given to the substance formed when haemoglobin in your red blood cells temporarily binds with oxygen.

Painkiller A drug that treats the symptoms of disease, such as a headache, but does not kill any pathogens that may be causing the disease.
Palisade mesophyll Tissue found towards the upper surface of leaves with lots of chloroplasts for photosynthesis.
Partially permeable Allowing only substances of a certain size through.
Pathogen A disease-causing microorganism (e.g. a bacterium or fungus).
Peer review A process by which scientists check each other’s work.
Period (in Chemistry) The name given to a row in the periodic table.
Period (in Physics) The time taken to produce one wave.
Phagocyte A type of white blood cell that engulfs pathogens.
Phloem Living cells that carry sugars made in photosynthesis to all cells of a plant.
Photosynthesis A chemical reaction that occurs in the chloroplasts of plants and algae and stores energy in glucose.
 Pipette A glass tube used to measure volumes of liquids with a very small margin of error.
Placebo A medicine that has only psychological effects.
Planet A large body which orbits the Sun.
Platelets Small structures (not cells) in your blood that fuse together to form a scab.
Polymer Long chain molecule made from joining lots of small molecules together.
Population The total number of all the organism of the same species or the same group of species that live in a particular geographical area.
Potential difference (p.d.) A measure of the electrical work done by a cell (or other power supply) as charge flows round the circuit; potential difference is measured in volts (V).
Power The rate at which energy is transferred.

\[
\text{Power} = \frac{\text{energy transferred}}{\text{time}}
\]
Sink  A long-term store of a substance, often carbon.
Soluble  Can dissolve.
Species  The smallest group of classifying organisms, all of which are able to interbreed to produce fertile offspring.
Specific heat capacity  The energy needed to raise the temperature of 1 kg of substance by 1 °C.
Spectator ions  Ions that do not take part in a reaction and do not appear in the ionic equation for the reaction.
Spongy mesophyll  Tissue found towards the bottom surface of leaves with spaces between the cells to allow gases to diffuse.
States of matter  These are solid, liquid and gas.
Statin  Drug that reduces blood cholesterol.
Stem cell  An undifferentiated cell that can develop into one or more types of specialised cell.
Stent  A small medical device made from mesh that keeps arteries open.
Step-down transformer  A transformer that decreases potential difference and increases current.
Step-up transformer  A transformer that increases potential difference and decreases current.
Strong acid  Acid in which all the molecules break into ions in water.
Substrate  The molecule or molecules on which an enzyme acts.
Systematic error  A consistent error, usually caused by the measuring instruments, when all of the data is higher or lower than the true value; data with a systematic error will give a graph line that is higher or lower than it should be.
Thermal decomposition  Reaction where high temperature causes a substance to break down into simpler substances.
Tissue fluid  The liquid that surrounds (‘bathes’) cells in the body tissues. It is formed from plasma that diffuses through the capillary walls.
Toxin  A poison that damages tissues and makes us feel ill.
Translocation  The movements of sugars made in photosynthesis from the leaves of plants.
Transparent  An object which allows us to see clearly through it; glass is transparent.
Transpiration  The gradual release of water vapour from leaves to continue the ‘pull’ of water up to them from the soil.
True value  The value that would be obtained in an ideal measurement.
Turgid  Describes swollen cells.
Type 2 diabetes  A medical condition that usually develops in later life, preventing cells from responding to insulin.
Uncertainty  The range of measurements within which the true value can be expected to lie.
Vaccine  A medicine containing an antigen from a pathogen that triggers a low level immune response so that subsequent infection is dealt with more effectively by the body’s own immune system.
Variation  The differences that exist within a species or between different species.
Vector  An organism that spreads a communicable disease.
Vein  A large blood vessel that returns blood to the heart.
Ventilation  Breathing in (inhaling) and out (exhaling).
Ventricle  A lower chamber of the heart surrounded by a thicker wall of muscle.
Villi (singular villus)  Tiny finger-like projections that increase the surface area of the small intestine.
Wavelength  The distance from a point on one wave to the equivalent point on the next wave.
Weak acid  Acid in which only a small fraction of the molecules break into ions in water.
Work  When a force causes an object to move. Work = force × distance
Xylem  Dead plant cells joined together into long tubes through which water flows during transpiration.
Yield  The amount of an agricultural product.
Zero error  When a measuring instrument gives a reading when the true value is zero.
1 becquerel (1Bq)  An emission of 1 particle/second.
1 GW, 1 gigawatt = 10⁹ W
### Index

#### A

- acid rain 277
- acids 209, 247
  - dilute and concentrated 211
  - reaction with alkalis 249, 252–3
  - reaction with ammonia 248
  - reaction with metal carbonates 214, 248
  - reaction with metal hydrogen carbonates 232
  - reaction with metal hydroxides 213–14, 248
  - reaction with metal oxides 214, 248, 249
  - reaction with metals 202, 205, 213, 248
  - strong and weak 210–11
- activation energy 233
- active sites 49
- active transport 36–7
- adaptations
  - cell specialisations 8–10
  - facilitation of diffusion 31, 33
- aerobic respiration 5, 105–6
  - conversion of energy 107
  - physical properties 162
- alkali metals 129–31
  - properties of 162
- alkalis 209, 247
  - reaction with acids 249, 252–3
- alloys 162
  - alpha particles 344, 345
  - properties of 346, 347
- alpha scattering 341–2
- of an irregularly shaped solid 321–2
- amperes (amps, A) 294
- amylase 44, 47
- anaerobic respiration 108–10
- anomalous results 322
- antibiotics 88–9
- antibodies 55, 86
- antigens 55, 86
- antitoxins 87
- anus 46
- anxiety 59
- aorta 52, 53
- arteries 51, 53
- atomic number 119, 339
- atomic structure 117–18, 338
  - Bohr model 342
  - development of ideas about 123–4
  - electronic structure 121
- atria 52, 53
- Avogadro constant 178–9
- bond energies 235–7
- bond energies 163
- giant covalent substances 159–60
- crystallisation 137, 138
- current–potential difference (I–V) graphs 298–9
  - investigating the characteristics of a circuit component 300
- data presentation 40
- density 319, 323
  - of an irregularly shaped solid 321–2
  - of a liquid 320
  - of a regular solid 321
- dependent variables 103, 241
- charge 294
- chemical energy 260, 282, 283
- chlorophyll 6, 96
- chloroplasts 6, 96
- chromatography 137, 139
- chromosomes 4, 19–20, 27
- cilia 85
- circuit calculations 304
- circuits
  - parallel 302
  - series 295, 301–2
- circuit symbols 293
- circulatory system 51
  - blood 54–6
  - blood vessels 52, 53–4
  - heart 51–3
  - clones 23, 24
- combustion reactions 232
- communicable diseases 79
  - bacterial 82
  - fungal 83
  - human defence systems 84–7
  - protist 84
  - spread of pathogens 80
  - viral 80–1
- compounds 126, 136–7
  - formulae for 244
- concentrated acids 211
- concentration gradients 29, 31, 32
- concentrations of solutions 193
- condensation 324, 328
- conservation of energy 260
- conservation of mass 181
- contamination by radioactive isotopes 351
- continuous variables 241
- control variables 103, 241
- cooling curves 329
- coronary heart disease 56–7
- coulombs (C) 294
- count rate, radioactive sources 348
- covalent bonds 154, 156–7, 163
- giant covalent substances 159–60
- current–potential difference (I–V) graphs 298–9
- data presentation 40
- density 319, 323
  - of an irregularly shaped solid 321–2
  - of a liquid 320
  - of a regular solid 321
- dependent variables 103, 241
Index

depression 59
diamond 166
diatomic molecules 244
diet, well-balanced 58–9
diffusion 5, 29
  influencing factors 32–3
  in the lungs 30
  in other organisms 31–2
digestive enzymes 46–7
digestive system 43–6
  active transport in 37
  model of 66
dilute acids 211
diodes 293, 298
diploid cells 19
direct current (d.c.) 305
diseases
  see communicable diseases; non-communicable diseases; pathogens
  displacement reactions 203, 205, 249
  half equations 206, 256–7
  ionic equations 206, 254
  distillation 137, 138–9
  DNA (deoxyribonucleic acid) 3, 19
drugs 88
  antibiotics 88–9
  discovery and development of 90–1
  painkillers 90
displacement reactions 203, 205, 249
  half equations 206, 256–7
  ionic equations 206, 254
  distillation 137, 138–9
  DNA (deoxyribonucleic acid) 3, 19
drugs 88
  antibiotics 88–9
  discovery and development of 90–1
  painkillers 90
deep reading 341

depression 59
diamond 166
diatomic molecules 244
diet, well-balanced 58–9
diffusion 5, 29
  influencing factors 32–3
  in the lungs 30
  in other organisms 31–2
digestive enzymes 46–7
digestive system 43–6
  active transport in 37
  model of 66
dilute acids 211
diodes 293, 298
diploid cells 19
direct current (d.c.) 305
diseases
  see communicable diseases; non-communicable diseases; pathogens
  displacement reactions 203, 205, 249
  half equations 206, 256–7
  ionic equations 206, 254
  distillation 137, 138–9
  DNA (deoxyribonucleic acid) 3, 19
drugs 88
  antibiotics 88–9
  discovery and development of 90–1
  painkillers 90
displacement reactions 203, 205, 249
  half equations 206, 256–7
  ionic equations 206, 254
  distillation 137, 138–9
  DNA (deoxyribonucleic acid) 3, 19
drugs 88
  antibiotics 88–9
  discovery and development of 90–1
  painkillers 90

eight wires 306

deep reading 341

depression 59
diamond 166
diatomic molecules 244
diet, well-balanced 58–9
diffusion 5, 29
  influencing factors 32–3
  in the lungs 30
  in other organisms 31–2
digestive enzymes 46–7
digestive system 43–6
  active transport in 37
  model of 66
dilute acids 211
diodes 293, 298
diploid cells 19
direct current (d.c.) 305
diseases
  see communicable diseases; non-communicable diseases; pathogens
  displacement reactions 203, 205, 249
  half equations 206, 256–7
  ionic equations 206, 254
  distillation 137, 138–9
  DNA (deoxyribonucleic acid) 3, 19
drugs 88
  antibiotics 88–9
  discovery and development of 90–1
  painkillers 90
displacement reactions 203, 205, 249
  half equations 206, 256–7
  ionic equations 206, 254
  distillation 137, 138–9
  DNA (deoxyribonucleic acid) 3, 19
drugs 88
  antibiotics 88–9
  discovery and development of 90–1
  painkillers 90

eight wires 306

deep reading 341

depression 59
diamond 166
diatomic molecules 244
diet, well-balanced 58–9
diffusion 5, 29
  influencing factors 32–3
  in the lungs 30
  in other organisms 31–2
digestive enzymes 46–7
digestive system 43–6
  active transport in 37
  model of 66
dilute acids 211
diodes 293, 298
diploid cells 19
direct current (d.c.) 305
diseases
  see communicable diseases; non-communicable diseases; pathogens
  displacement reactions 203, 205, 249
  half equations 206, 256–7
  ionic equations 206, 254
  distillation 137, 138–9
  DNA (deoxyribonucleic acid) 3, 19
drugs 88
  antibiotics 88–9
  discovery and development of 90–1
  painkillers 90
displacement reactions 203, 205, 249
  half equations 206, 256–7
  ionic equations 206, 254
  distillation 137, 138–9
  DNA (deoxyribonucleic acid) 3, 19
drugs 88
  antibiotics 88–9
  discovery and development of 90–1
  painkillers 90

eight wires 306

deep reading 341
I
independent variables 103, 241
inert electrodes 221
infectious diseases 79
internal (thermal) energy 259, 260, 282, 324
changes in 326
in vitro fertilisation (IVF) 24
ionic bonding 149, 163
ionic equations 252
for displacement reactions 206, 254
for the reaction of acids with alkalis 252–3
ionic substances 148, 163–4, 247
electrolysis of 219–20, 221–3
formation of 152–3
formulae of 150–1, 245–6
properties of 149–50, 159
ionisation 345
ions 121–2, 148, 339
charge on 150–1
irradiation 351
isotonic solutions 34
isotopes 119–20, 340

J
joules (J) 266

K
kinetic energy 259, 260
calculation of 262, 263

L
lactic acid 108–9
large intestine 45–6
latent heat 327–8
leaves 71
light, models of 336
light-dependent resistors (LDRs) 293, 299–300
light-emitting diodes (LEDs) 293, 298
light microscopes 11
limiting reagents 190–1
lipids, test for 48
liquids 164–5, 323
density of 320, 323
liver 45, 47
live wires 306
lock and key hypothesis 49–50, 65
lungs
adaptations of 31
gas exchange 30
lymphocytes 55, 86

M
machines 275
magnification 15–17
mains electricity 305–6
malaria 84
malleability 161
mass
conservation of 181
units of 193–4
mass number 119, 339
means 115
measles 80–1, 94
measurements 228–9
melting 324, 328
melting and boiling points 164–5
of giant covalent substances 160
of ionic substances 149
of metallic substances 161
of molecular substances 154–5
Mendeleev, Dmitri 135–6
meniscus 228
mental health 59
meristems 23, 69, 70
metabolism 110–11
metal carbonates, reaction with acids 214, 248
metal hydrogencarbonates, reaction with acids 232
metal hydroxides, reaction with acids 213–14, 248
metallic bonding 160–1, 163
metal oxides
acid–base character of 247, 248
reaction with acids 214, 248, 249
metals 163–4, 247
alloys 162
displacement reactions 203
extraction from ore 208, 220–1
and non-metals 125
properties of 161
reaction with acids 248
reaction with dilute acids 202, 205, 213
reaction with non-metals 152–3, 249
reaction with oxygen 181, 182, 201–2, 205
reaction with water 202, 205, 248
reactivity series of 200–1
structure of 160–1
what happens when they react 203
microscopy
electron microscopes 12
examination of cells 7–8
light microscopes 11
magnification 15–17
mitochondria 5, 105
mitosis 20–1
mixtures 136–7
separation of 137–40
MMR vaccination 87, 94
models 65–6, 335–6
of atomic structure 123–4, 342
of radioactive decay 347–8
molar ratios 185–6
molecular substances 154, 163–4, 247
drawing molecules 157
formulae of 155–6
properties of 154–5, 159
moles 178–9
monatomic substances 163–4, 247
monomers 158
muscle cells 9

N
nanotubes, carbon 168–9
National Grid 308–9
nerve cells 9
neutralisation reactions 214, 232
temperature changes 234
neutral wires 306
neutron emission 344
neutrons 117–18, 338, 339
discovery of 342
noble gases 128–9
non-communicable diseases
effect of lifestyle 61–2
heart disease 56–8
non-metals 125
reaction with metals 152–3, 249
non-renewable energy resources 276
nuclear decay 343–4
nuclear power 278
nucleus of a cell 5
nucleus of an atom 117–18, 338
discovery of 341–2

O
oesophagus 44
ohmic conductors 298
OIL RIG phrase 205
ores 208
metal extraction 220–1
organelles 4
organisational levels 43
osmosis 33–5
investigation of 41
oxidation reactions 105, 201, 232
in terms of electrons 204–5
oxides
acid–base character of 247–8
see also metal oxides
oxygen, reaction with metals 181, 182, 201–2, 205
oxygen debt 108–9
oxyhaemoglobin 55

P
pacemaker 51
problems with 57, 58
painkillers 90
palisade mesophyll 68, 69
pancreas 45, 47
parallel circuits 302
partially permeable membranes 33–4
Index

particle model of gases 330–1, 335–6
pathogens
defence against 84–8
spread of 80
periodic table 125, 357
and electronic structure 127–8
Group 0 (noble gases) 128–9
Group 1 (alkali metals) 129–31, 162
Group 7 (halogens) 131–4
history of 134–6
peristalsis 45–6
phagocytes 55, 86–7
phloem 10, 69
photosynthesis 6
chemical reaction of 96–7
conversion of energy 107
effect of light intensity 99
limiting factors 97–8
source of energy for 100
uses of glucose produced 100
pH scale 209–10
plant cells 6, 9–10, 58
plant organs 70–1
plant tissues 68–9
plant transport organ system 71–3
plasma 54, 55, 56
plasmids 3, 4
platelets 56
polymers 158
potential difference (p.d.) 294
power 266–7
of electrical appliances 307–8
powers of ten 316
power stations 272, 277
hydroelectric 279
nuclear 278
precision 289
predictions 27
pressure, of gases 330–1
prokaryotes 2–4
proportion, direct and inverse 297
protein, test for 48
protist diseases 84
protons 117–18, 338, 339
pulley systems 275
pulmonary artery 52, 53
pulmonary vein 52, 54

R
radiation 282
radioactive decay 343–4, 347–9
radioactivity
detection of 346, 348
discovery of 343
effects on the body 350
ionisation 345
irradiation and contamination 351
properties of 346–7
random errors 76, 289
range of data 115
reacting mass calculations 186–7
reaction mechanisms 249
reaction profiles 232–3, 235
reactions between elements 126
reactivity series of metals 200–1
reactivity trends
alkali metals 131
halogens 132–3
red blood cells 54–5
redox reactions 205
reduction reactions 201
in terms of electrons 204–5
relative atomic mass 120, 177
relative formula mass 177–8
renewable energy resources 276
repeatability of data 115
relationship to the length of a wire 297
resistors 293, 295–6
light-dependent (LDRs) 299–300
ohmic and non-ohmic 298
thermistors 298–9
resolution of apparatus 228
respiration
aerobic 5, 105–7
anaerobic 108–10
ribosomes 4, 5
risk assessment 356
risk factors for disease 61–2
rock salt extraction 140
root hair cells 9
active transport 36
roots 70
rose black spot 83
salivary glands 44, 47
salmonella 82
salts 247
making soluble salts 216–18
what they are 216
scientific thinking 94
screening 60
separating tunnels 137, 139
series circuits 295, 301–2
shoots 71
significant figures 180
SI units 175
interconversion of 193–4
skin 85
small intestine 45
solar power 280
solids 164–5, 323
density of 321–2, 323
solutions, concentrations of 193
specific heat capacity 268, 325–6
in investment of 269–70
of water 326

working out the temperature of a hot object 270–1
specific latent heat 327–8
spectator ions 252
sperm cells 8
spongy mesophyll 68, 69
standard form 174–5
starch, test for 48
states of matter 164–5, 323
changes of 324–5
state symbols 184, 251
stem cells 22–4
step-up and step-down transformers 309
stomach 44, 47
acid production 85
stomata 68–9
stress 59
strong acids 210–11
structure of substances 247
sublimation 324
surface area, effect on diffusion 32, 33
switches 293
synthesis enzymes 50
synthesis reactions 110–11
systematic errors 76, 289

T
tables 40
temperature, changes during
neutralisation reactions 234
temperature scales 316
theories, changes over time 145–6
thermal (internal) energy 259, 260, 282, 324
changes in 326
thermal conductivity, of metallic substances 161
thermal decomposition 182, 232
of sodium hydrogen carbonate 188–9
thermal insulation 272–3
thermistors 293, 298–9, 316–17
thermosoftening polymers 158
Thompson, J.J. 341
tidal energy 278
tobacco mosaic virus 81
transformers 309
translocation 10, 72
transpiration 72
investigation of 73

U
uncertainties 228–9
universal indicator 210

V
vaccination 87–8, 94
valves of the heart 53
problems with 57, 58
Index

vaporising 328
variable resistors 293, 296
variables 103, 241–2
veins 51, 53–4
vena cava 52, 54
ventricles 52, 53
villi 45
viral diseases 80–1
voltmeters 293, 296
 calibration of 316–17
volts (V) 294
volume, units of 193

W
water
 latent heat of vaporisation 328
 properties of 155
 reaction with metals 130, 131, 205, 248
 specific heat capacity 326
 water concentrations 34
 watts (W) 267
 weak acids 210–11
 white blood cells 55
 wind power 279
 work 265–6

X
xylem 10, 69

Z
zero errors 289