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- Track student’s progress

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- Suggested Schemes of work for 2- and 3-year teaching
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- Add notes and highlight areas
- Add double-page spreads into lesson plans

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- Add, edit and synchronize notes across two devices
- Access their personal copy on the move

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<table>
<thead>
<tr>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
<th>7</th>
<th>8</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>hydrogen</td>
<td>He</td>
<td>helium</td>
<td>Li</td>
<td>lithium</td>
<td>Be</td>
<td>beryllium</td>
</tr>
<tr>
<td>Na</td>
<td>sodium</td>
<td>Mg</td>
<td>magnesium</td>
<td>Al</td>
<td>aluminium</td>
<td>Si</td>
<td>silicon</td>
</tr>
<tr>
<td>K</td>
<td>potassium</td>
<td>Ca</td>
<td>calcium</td>
<td>Ti</td>
<td>titanium</td>
<td>V</td>
<td>vanadium</td>
</tr>
<tr>
<td>Rb</td>
<td>rubidium</td>
<td>Sr</td>
<td>strontium</td>
<td>Zr</td>
<td>zirconium</td>
<td>Nb</td>
<td>niobium</td>
</tr>
<tr>
<td>Cs</td>
<td>cesium</td>
<td>Ba</td>
<td>barium</td>
<td>La*</td>
<td>lanthanum</td>
<td>Ce</td>
<td>cerium</td>
</tr>
<tr>
<td>Fr</td>
<td>francium</td>
<td>Ra</td>
<td>radium</td>
<td>Ac*</td>
<td>actinium</td>
<td>Th</td>
<td>thorium</td>
</tr>
</tbody>
</table>

* 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.
Get the most from this book

Welcome to the AQA GCSE Chemistry Student Book.
This book covers the Foundation and Higher-tier content for the 2016
AQA GCSE Chemistry specification.
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.

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.

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

Required practical
AQA required practicals are clearly highlighted.

TIPS
These highlight important facts, common misconceptions and signs 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.

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

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.

Answers
Answers for all questions and activities in this book can be found online at www.hoddereducation.co.uk/aqagceschemistry

Practice questions
You will find Practice questions at the end of every chapter. These follow the style 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 activity at the end of each chapter. Work through these activities on your own or in groups. You will develop skills such as Dealing with data, Scientific thinking and Experimental skills.
Previously you could have learned:
> Elements are made of particles called atoms.
> Elements are substances containing only one type of atom – this means they cannot be broken down into simpler substances.
> Each element has its own symbol and is listed in the periodic table.
> Elements are either metals or non-metals.
> Compounds are substances made from atoms of different elements bonded together.
> Compounds have different properties from the elements from which they are made.
> Compounds are difficult to break back down into their elements.
> Substances in mixtures are not chemically joined to each other.
> Substances in mixtures can be separated easily by a range of techniques.

Test yourself on prior knowledge
1. What is an element?
2. What is a compound?
3. Why do compounds have different properties from the elements from which they are made?
4. List some differences between metals and non-metals.
5. Why is it easy to separate the substances in a mixture but not to break apart a compound?
6. Name four methods of separating mixtures.

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 1.1).

<table>
<thead>
<tr>
<th>Particle</th>
<th>Relative Mass</th>
<th>Relative Charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>1</td>
<td>+1</td>
</tr>
<tr>
<td>Neutron</td>
<td>1</td>
<td>0</td>
</tr>
<tr>
<td>Electron</td>
<td></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 (0.000000001 m, that is 1 x 10^-10 m). Atoms have a central nucleus which contains protons and neutrons (Figure 1.1). 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 1/100000 of that of the atom (1 x 10^{-10} 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 1.2).

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 0.077 nm. Write this in standard form in the units of metres.
2. The radius of a hydrogen atom is 2.5 x 10^{-10} m. Write this in nanometres.
3. The radius of a chlorine atom is 1.9 x 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 0.074 nm. The nucleus of an atom is about 100 000 times smaller. Estimate the radius of the nucleus of a sodium atom. Verify your answer in both nanometres and metres.
5. A copper atom has a diameter of 0.125 nm. A copper wire has a diameter of 0.00440 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 0.176 nm. The largest gold bar in the world is 6.5 cm long. How many gold atoms fit into it? 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.

**Example**

How many protons, neutrons and electrons are there in an atom of Pb?"
This can be calculated as shown:

\[ \text{Relative atomic mass} \times \text{total number of atoms} = \text{total mass of all atoms of element} \]

**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**

Relative atomic mass \( = \frac{(0.75 \times 35) + (0.25 \times 37)}{1} = 35.5 \)

---

**Test yourself**

1. List the three particles found inside atoms.
2. Identify the particle found inside the nucleus of an atom that has no charge.
3. Atoms contain positive and negative particles. Explain why atoms are neutral.
4. How many protons, neutrons and electrons are there in an atom of 14P?
5. What is it about the atom 83Kr that makes it an atom of potassium?

12. Describe the similarities and differences between atoms of the isotope 83Kr and 92Kr.
13. The element copper contains 65\% Cu and 35\% Cu. Find the relative atomic mass of copper to one decimal place. Show your working.
14. The element magnesium contains 79\% Mg, 10\% Mg and 11\% Mg. Find the relative atomic mass of magnesium to one decimal place. Show your working.
15. Explain why mass number is an integer, but relative atomic mass is not.

---

**Show you can...**

Copy and complete the table for each of the elements below:

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic number</th>
<th>Mass number</th>
<th>Number of protons</th>
<th>Number of electrons</th>
<th>Number of neutrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>p</td>
<td>14</td>
<td>32</td>
<td>14</td>
<td>14</td>
<td>18</td>
</tr>
<tr>
<td>14S</td>
<td>16</td>
<td>32</td>
<td>16</td>
<td>16</td>
<td>16</td>
</tr>
<tr>
<td>16S</td>
<td>16</td>
<td>36</td>
<td>16</td>
<td>16</td>
<td>20</td>
</tr>
</tbody>
</table>

---

**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 occupying the third energy level (third shell). The next two electrons occupy the fourth energy level (fourth shell). The arrangement of electrons in some atoms are shown in Table 1.3. The electron structure can be drawn on a diagram or written using numbers. For example, the electron 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 1.3).

---

**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 +1\text{+}) and 10 electrons (total charge 1\text{+}) will have an overall charge of +1\text{+}.
- An ion with 16 protons (total charge 16\text{+}) and 18 electrons (total charge 18\text{+}) will have an overall charge of 2\text{+}.

Table 1.4 shows some common ions.

**TIP**

It is usual to write ion charges with the number before the + or – sign, such as +2\text{+}, but it is not wrong to write it as +2.

**KEY TERM**

Ion - An electrically charged particle containing different numbers of protons and electrons.

---

**Simple ions (those made from single atoms)** have the same electron structure as the elements in Group 8 of the periodic table (Table 1.5). The elements in Group 8 are called the noble gases. The noble gases have very stable electron structures.

**TIP**

Positive ions have more protons than electrons. Negative ions have more electrons than protons. This is because electrons are negatively charged.
The hydrogen ion (H\(^+\)) is the only simple ion that does not have the electron 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.

**Test yourself**

10 What is the charge of a particle with 19 protons and 18 electrons?
11 What is the charge of a particle with 7 protons and 10 electrons?
20 What is the electron structure of the F\(^-\) ion?
21 How many protons, neutrons and electrons are there in the F\(^-\) ion?
22 What is the link between the electron structure of ions and the group 8 elements (the noble gases)?

Show you can...

Table 1.4 gives some information about six different particles, A, B, C, D, E and F. Some particles are atoms and some are ions. The letters are not chemical symbols.

<table>
<thead>
<tr>
<th>Particle</th>
<th>Protons</th>
<th>Neutrons</th>
<th>Number of protons</th>
<th>Number of neutrons</th>
<th>Number of electrons</th>
<th>Electron structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>18</td>
<td>8</td>
<td>18</td>
<td>8</td>
<td>20</td>
<td>2,8</td>
</tr>
<tr>
<td>B</td>
<td>17</td>
<td>8</td>
<td>17</td>
<td>8</td>
<td>20</td>
<td>2,8</td>
</tr>
<tr>
<td>C</td>
<td>16</td>
<td>8</td>
<td>16</td>
<td>8</td>
<td>18</td>
<td>2,8,7</td>
</tr>
<tr>
<td>D</td>
<td>17</td>
<td>8</td>
<td>17</td>
<td>8</td>
<td>20</td>
<td>2,8</td>
</tr>
<tr>
<td>E</td>
<td>16</td>
<td>8</td>
<td>16</td>
<td>8</td>
<td>18</td>
<td>2,8,7</td>
</tr>
<tr>
<td>F</td>
<td>17</td>
<td>8</td>
<td>17</td>
<td>8</td>
<td>20</td>
<td>2,8</td>
</tr>
</tbody>
</table>

a) Copy and complete the table.
b) Particle C is an atom. Explain, using the information in the table, why particle C is an atom.
c) Particle D is a negative ion. What is the charge on this ion?
d) Which two atoms are isotopes of the same element?

**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 atoms 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 1.4).

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 1.5).

In 1913, Niels 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 1.6).

The development of ideas about atomic structure shows how well 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 up to just over 100. This means that there are just over 100 elements. All the elements are listed in the periodic table. The elements are listed in order of atomic number (Figure 1.7).

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.

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 1.7, although there are some exceptions.

<table>
<thead>
<tr>
<th>Property</th>
<th>Metals</th>
<th>Non-metals</th>
</tr>
</thead>
<tbody>
<tr>
<td>Melting and boiling points</td>
<td>High</td>
<td>Low</td>
</tr>
<tr>
<td>Conductivity</td>
<td>Thermal conductor</td>
<td>Thermal and electrical insulator except graphite</td>
</tr>
<tr>
<td>Density</td>
<td>High density</td>
<td>Low density</td>
</tr>
<tr>
<td>Reactivity</td>
<td>Shiny when polished</td>
<td>Dull</td>
</tr>
<tr>
<td>Malleability</td>
<td>Can be hammered into shape</td>
<td>Entire as solids</td>
</tr>
<tr>
<td>Reactivity with non-metals</td>
<td>React to form positive ions in ionic compounds</td>
<td>React to form negative ions in ionic compounds</td>
</tr>
<tr>
<td>Reactivity with metals</td>
<td>Non-reactive</td>
<td>React to form compounds made of molecules</td>
</tr>
<tr>
<td>Acid-base properties of oxides</td>
<td>Metal oxides are basic</td>
<td>Non-metal oxides are acidic</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

2b Is each of the following elements a metal or non-metal?
   a. Element 1 is a dull solid at room temperature that readily 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.
   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 1.8 shows magnesium and oxygen reacting to form a white 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 up of different elements bonded 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 electron structure of the noble gases (Group 8 elements).

Table 1.9 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 electron 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>Molecular compound</td>
<td>Hydrogen reacts with oxygen by sharing electrons and forming molecules of water</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>Non-metal + Metal</td>
<td>Electrons transferred from metal to non-metal</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>
<td></td>
</tr>
<tr>
<td>Non-metal + Metal</td>
<td>No reaction as both metals cannot lose electrons</td>
<td>Covalent bond</td>
<td>Molecular compound</td>
<td>Hydrogen reacts with oxygen by sharing electrons and forming molecules of water</td>
</tr>
</tbody>
</table>

**Show you can...**

The electron 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: 2,8  
D: 2,8,7  
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.

**Test yourself**

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

- potassium + oxygen
- bromine + iodide
- oxygen + sulfur
- sulfur + magnesium
- calcium + potassium
- nitrogen + hydrogen

**TIP**

The chemical properties of the elements in the periodic table repeat at regular (periodic) intervals. This is why it is called the periodic table.
### Group 0 – the noble gases

The main elements of Group 0 are helium, neon, argon, krypton, xenon and radon (Figure 1.11 and Table 1.9). They are known as the noble gases (Table 1.10). These atoms all have stable electron structures. Helium’s outer shell is full with 2 electrons while the others have 8 electrons in their outer shells.

<table>
<thead>
<tr>
<th>Element</th>
<th>Formula</th>
<th>Appearance at room temperature</th>
<th>Number of electrons in outer shell</th>
<th>Atomic mass of atom (g/mol)</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>3.60</td>
<td>45</td>
</tr>
<tr>
<td>Neon</td>
<td>Ne</td>
<td>Colourless gas</td>
<td>8</td>
<td>20.18</td>
<td>85</td>
</tr>
<tr>
<td>Argon</td>
<td>Ar</td>
<td>Colourless gas</td>
<td>8</td>
<td>39.94</td>
<td>151</td>
</tr>
<tr>
<td>Krypton</td>
<td>Kr</td>
<td>Colourless gas</td>
<td>8</td>
<td>83.80</td>
<td>112</td>
</tr>
<tr>
<td>Xenon</td>
<td>Xe</td>
<td>Colourless gas</td>
<td>8</td>
<td>131.30</td>
<td>115</td>
</tr>
<tr>
<td>Radon</td>
<td>Rn</td>
<td>Colourless gas</td>
<td>8</td>
<td>222.00</td>
<td>69</td>
</tr>
</tbody>
</table>

### Test yourself

28 In what order are the elements in the periodic table?
29 In which group of the periodic table do the elements with these electron structures belong?
30 Explain why the periodic table has the words periodic in its name.

### Group 1 – the alkali metals

The main group of Group 1 are lithium, sodium, potassium, rubidium and caesium (Table 1.11). They are known as the alkali metals (Table 1.12). 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 1.12). They are very reactive and so are stored in bottles of oil to stop them reacting with water and oxygen.

### Table 1.11: Properties of the alkali metals

<table>
<thead>
<tr>
<th>Element</th>
<th>Formula</th>
<th>Melting point</th>
<th>Boiling point</th>
<th>Density at 20°C</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lithium</td>
<td>Li</td>
<td>196°C</td>
<td>1508°C</td>
<td>0.53 g/cm³</td>
</tr>
<tr>
<td>Sodium</td>
<td>Na</td>
<td>97°C</td>
<td>982°C</td>
<td>0.97 g/cm³</td>
</tr>
<tr>
<td>Potassium</td>
<td>K</td>
<td>63°C</td>
<td>762°C</td>
<td>0.89 g/cm³</td>
</tr>
<tr>
<td>Rubidium</td>
<td>Rb</td>
<td>39°C</td>
<td>672°C</td>
<td>1.53 g/cm³</td>
</tr>
<tr>
<td>Caesium</td>
<td>Cs</td>
<td>28°C</td>
<td>808°C</td>
<td>1.93 g/cm³</td>
</tr>
</tbody>
</table>

### Test yourself

31 Why are the noble gases unreactive?
32 Suggest a reason why the noble gases are not found as being in Group 0 rather than Group 0.
33 Some atoms of element 118 (Og) have been produced. Element 118 is in Group 0. Predict the chemical and physical properties of this element.
Reactivity trend of the alkali metals
The alkali metals get more reactive as you go down the group (Figure 1.13). This can be seen when the alkali metals react with water.

<table>
<thead>
<tr>
<th>Description</th>
<th>Lithium (Li)</th>
<th>Potassium (K)</th>
<th>Caesium (Cs)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Reacts with water</td>
<td>Forms amalgam</td>
<td>Forms very reactive solution</td>
<td>Forms very reactive solution</td>
</tr>
<tr>
<td>Name of product</td>
<td>Lithium amalgam</td>
<td>Potassium hydroxide and hydrogen</td>
<td>Caesium hydroxide and hydrogen</td>
</tr>
</tbody>
</table>

When the alkali metals react with water, they lose their outer shell electron in order to get a noble gas electron 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 it is to lose, the more reactive the alkali metal.

Test yourself

1. Why are the alkali metals reactive?
2. Write word and balanced symbol equations for the reaction of potassium with water.
3. Explain why the solution formed when potassium reacts with water has a high pH.
4. Potassium reacts with chlorine to form an ionic compound. Explain why this reaction happens.
5. Explain why potassium is more reactive than sodium.
6. Francium is the last element in Group 1. Predict the chemical and physical properties of francium.

Show you can...

This question gives information about the reactions of Group 1 elements with water (these reactions are not on the specification) and tests your ability to interpret data.

- **Reactivity with water**
  - **Lithium**
  - **Potassium**
  - **Caesium**

- **Name of product**
  - Lithium amalgam
  - Potassium hydroxide and hydrogen
  - Caesium hydroxide and hydrogen

Use the information in the table and your own knowledge of Group 1 elements to compare and contrast the reactions of Group 1 and Group 2 elements with water.

- The products formed.
- The reactivity of the Group 1 elements compared to the Group 2 elements.
- The trend in reactivity down both groups.

Group 7 – the halogens
The main elements of Group 7 are fluorine, chlorine, bromine, and iodine (Figure 1.14) and these are known as the halogens (Table 1.14). 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 electron structure. The particles in each of the elements are molecules containing two atoms (diatomic molecules), such as F₂, Cl₂, Br₂, and I₂.

### Table 1.13

<table>
<thead>
<tr>
<th>Element</th>
<th>Formula</th>
<th>Appearance</th>
<th>Number of electrons in outer shell</th>
<th>Melting point °C</th>
<th>Boiling point °C</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fluorine</td>
<td>F₂</td>
<td>Light yellow gas</td>
<td>7</td>
<td>182</td>
<td>250</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl₂</td>
<td>Pale green gas</td>
<td>7</td>
<td>180</td>
<td>340</td>
</tr>
<tr>
<td>Bromine</td>
<td>Br₂</td>
<td>Dark brown liquid</td>
<td>7</td>
<td>72</td>
<td>590</td>
</tr>
<tr>
<td>Iodine</td>
<td>I₂</td>
<td>Grey solid</td>
<td>7</td>
<td>195</td>
<td>451</td>
</tr>
</tbody>
</table>

### Table 1.14 Properties of the halogens

- **Reactivity**
  - Each of the halogens is reactive.
  - The halogens are the most reactive of the elements and they are used in many different applications.
  - The halogens react with other non-metals by forming compounds to form compounds to form compounds to form compounds to form compounds.
  - The halogens all react easily with metals by forming compounds made of ions.

Reactivity trend of the halogens

The halogens get less reactive as you go down the group. This can be seen by looking at which halogen can displace each other from compounds. Compounds containing halogen, such as sodium chloride and potassium bromide, are often called halides or halite 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 + Iron oxide → Aluminium oxide + 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 1.15 and Figure 1.15).

<table>
<thead>
<tr>
<th>Potassium chloride</th>
<th>Potassium bromide</th>
<th>Potassium iodide</th>
</tr>
</thead>
<tbody>
<tr>
<td>Chlorine (aq)</td>
<td>Sodium (aq)</td>
<td>Sodium (aq)</td>
</tr>
<tr>
<td>No reaction</td>
<td>Bromine cannot</td>
<td>Iodine cannot</td>
</tr>
<tr>
<td></td>
<td>displace chlorine</td>
<td>displace bromine</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Chlorine + potassium bromide → potassium chloride + bromine</td>
<td>C(_2)H(_5)I + 2Cl(^-) → 2C(_2)H(_5)Cl + I(^-)</td>
<td>C(_2)H(_5)I + 2Cl(^-) → 2C(_2)H(_5)Cl + I(^-)</td>
</tr>
<tr>
<td>Yellow solution formed due to production of bromine</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>Chlorine displaces bromine</td>
<td></td>
</tr>
<tr>
<td>Chlorine + potassium iodide → potassium chloride + iodine</td>
<td>C(_2)H(_5)I + 2Cl(^-) → 2C(_2)H(_5)Cl + I(^-)</td>
<td>C(_2)H(_5)I + 2Cl(^-) → 2C(_2)H(_5)Cl + I(^-)</td>
</tr>
<tr>
<td>Brown solution formed due to production of iodine</td>
<td></td>
<td></td>
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<tr>
<td></td>
<td>Chlorine displaces iodine</td>
<td></td>
</tr>
<tr>
<td>Boron + potassium iodide → potassium bromide + iodine</td>
<td>B(_2)I(_3) + 2Br(^-) → 2B(_2)Br(_3) + I(^-)</td>
<td>B(_2)I(_3) + 2Br(^-) → 2B(_2)Br(_3) + I(^-)</td>
</tr>
<tr>
<td>Brown solution formed due to production of bromine</td>
<td></td>
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<tr>
<td></td>
<td>Chlorine displaces iodine</td>
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</tbody>
</table>
### Transition Metals

Transition metals have some properties in common with the alkali metals, but many differences (Table 1.16).

<table>
<thead>
<tr>
<th>Properties</th>
<th>Group I – Alkali Metals</th>
<th>Transition Metals</th>
</tr>
</thead>
<tbody>
<tr>
<td>Physical properties</td>
<td>Soft, low melting points, low density</td>
<td>Hard, high melting points, high density</td>
</tr>
<tr>
<td>Reactions</td>
<td>React with water or oxygen to form hydroxides or oxides</td>
<td>React with non-transition metals to form ionic compounds</td>
</tr>
</tbody>
</table>

### Test Yourself

- **Q1:** What are the transition metals?
- **Q2:** List some ways in which the transition metals are similar to the alkali metals.
- **Q3:** List some ways in which the transition metals are different from the alkali metals.
- **Q4:** What is a catalyst?

### 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 did lithium, sodium and potassium in the same group, each of which has very similar properties (Figure 2.20).
Over the next few years, elements were discovered that Mendeleev had predicted would exist (Table 1.17). These included gallium (1875), scandium (1879) and germanium (1886). In each case the properties of the element closely matched Mendeleev’s predictions. Table 1.17 shows some of the properties that Mendeleev predicted for the element he called eka-silicon and that we call germanium.

<table>
<thead>
<tr>
<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></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Oka-silicon</td>
<td>Grey metal</td>
<td>72</td>
<td>High</td>
<td>5.5</td>
<td>SiO₂</td>
<td>SiCl₄</td>
</tr>
<tr>
<td>Actual properties</td>
<td>Germainium</td>
<td>Grey-white metal</td>
<td>73</td>
<td>947</td>
<td>5.4</td>
<td>GeO₂</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.

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.

**Show you can...**

State some features of the periodic table developed by Mendeleev which are different from today’s modern periodic table.

**Test yourself**

49. In what order did Mendeleev put the elements?
50. Why did Mendeleev not stick to this order for some elements?
51. Why did Mendeleev leave some gaps in his table?
52. Why did Mendeleev’s ideas become accepted?

---

**Mixtures compared to compounds**

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 1.18).

---

**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 1.19).

---

**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.

---

**Salt**

Grey, highly reactive, dangerous metal
Green, highly reactive, toxic gas
White, non-toxic solid that we eat
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 1.22).

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 1.23).

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 1.24).

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. 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 vapourised solvent passes through a water-cooled condenser where it cools and condenses. The condensed solvent drips into a container away from the original solution (Figure 1.25).

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 (see Chapter 7), the whole mixture is vapourised 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 1.26).

Separating funnel
Liquids that do not mix together are called immiscible liquids. A hydrocarbon and water is an example of liquids that are immiscible with each other. They can be separated in a separating funnel. The funnel forms 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 1.27).

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 it, 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 1.28).

Chromatography is studied further in Chapter 8.

Test yourself
3. 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.
**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.

<table>
<thead>
<tr>
<th>Type of mixture separated</th>
<th>Important word and definition</th>
</tr>
</thead>
<tbody>
<tr>
<td>Filtration</td>
<td>Important word and definition</td>
</tr>
<tr>
<td>Distillation</td>
<td>Important word and definition</td>
</tr>
<tr>
<td>Fractional distillation</td>
<td>Important word and definition</td>
</tr>
</tbody>
</table>

**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.

1. Place 8 volumes of rock salt into a mortar and grind using a pestle.
2. Place the rock salt into a beaker and quarter fill with water.
3. Place a gas bottle and heat, stirring with a glass rod. Stop heating when the salt has dissolved - the sand and clay will be left undissolved.
4. Allow to cool and then filter.
5. Heat until half the volume of liquid is left.
6. Place the evaporation basin on the windscreen to evaporate off the rest of the water slowly. Pure salt crystals should be left.

**Chapter review questions**

1. Choose from the following list of elements to answer the questions below:
   - bromine
   - calcium
   - krypton
   - nickel
   - nitrogen
   - potassium
   - silicon
   - Which element is most like lithium?
   - Which element is most like iron?
   - Which element is most like helium?
   - Which element is most like fluorine?
   - Which element is most like carbon?
2. In which group or area of the periodic table would you find these elements?
   - Element A has 7 electrons in its outer shell.
   - Element B reacts vigorously with water to give off hydrogen gas and an alkaline solution.
   - Element C is a metal with 4 electrons in its outer shell.
   - Element D is a colourless gas that does not react at all.
   - Element E forms coloured compounds.
   - Element F is toxic and is made of diatomic molecules.
   - Element G forms 1 ion when it reacts with metals to form ionic compounds.
   - Element H can form both 1+ and 2+ ions.
   - Element I is a metal that floats on water.
   - Element J is the electron structure 2, 8, 18, 6.
   - Element K has 12 protons.
   - Element L has a full outer shell.
   - Element M can act as a catalyst.
3. Identify a mixture that could be separated by each of the following methods.
   - Simple distillation
   - Filtration
   - Chromatography
   - Crystallisation
   - Fractional distillation
4. Look at the following atoms and ions.
   - \( ^{12}_{4}C \), \( ^{9}_{4}Be \), \( ^{20}_{8}F \), \( ^{20}_{8}Ne \)
   - Which of these atoms and ions, if any:
     - are isotopes?
     - have 9 protons?
     - have 10 electrons?
     - have 10 neutrons?
     - have more protons than electrons?
5. Caesium atoms are among the largest atoms. A caesium atom has a radius of 0.260 nm. Write this in metres in standard form.
practice questions

1. How many electrons are there in a potassium ion ([K⁺]?)
   a. 18
   b. 20
   c. 19
   d. 18
   e. 20
   f. 19

2. In which of the following elements is the number of protons greater than the number of electrons?
   a. Ne
   b. N
   c. P
   d. Na
   e. Mg

3. What is the electron configuration of the element with 10 electrons?
   a. [H] 1
   b. [H] 1 1
   c. [H] 1 1 1
   d. [H] 1 1 1 1
   e. [H] 1 1 1 1 1

4. Write an ionic equation for the reaction between sodium and water.

5. Write a balanced equation for the reaction between sodium and bromine.

6. A yellow solution of bromine water was added dropwise to a colourless solution of sodium carbonate. The solution turned pink brown.
   a. Describe the reaction observed.
   b. Write an ionic equation for the reaction that took place.

7. The following four substances are mixed together: salt, water, cyclohexane, and cyclobutane. Explain why the reaction took place. Is sodium in water but not in cyclohexane or cyclobutane, then sodium is not reactant. Water is not reactant with cyclohexane or cyclobutane. Describe how the four substances could be separated.

8. A positive ion has 17 protons and 18 electrons. Is it an isotope of chlorine? Explain your answer.

9. What is the electron configuration of an element with 10 electrons?
   a. [H] 1
   b. [H] 1 1
   c. [H] 1 1 1
   d. [H] 1 1 1 1
   e. [H] 1 1 1 1 1

10. Write an ionic equation for the reaction between sodium and water.
The filter paper was placed in a test tube containing 10 drops of saline. The solution soaked up the paper and carried different chemicals. After 10 minutes, the filter paper was removed and allowed to dry. The results are shown:

a) What is the reason for the result shown by the student to analyze the two orange drops?

b) What is the reason for the result shown by the student to analyze the two orange drops?

c) What other factor could also cause the results to be different?

d) Complete the table below with the correct information.

<table>
<thead>
<tr>
<th>Group</th>
<th>Number of elements</th>
<th>Characteristics</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>2</td>
<td>Metallic</td>
</tr>
<tr>
<td>2</td>
<td>8</td>
<td>Non-metallic</td>
</tr>
<tr>
<td>3</td>
<td>18</td>
<td>Non-metallic</td>
</tr>
<tr>
<td>4</td>
<td>20</td>
<td>Non-metallic</td>
</tr>
</tbody>
</table>

Practice questions

1. What happens when the reactants in the Group 1 elements form a compound? (3 marks)

2. Explain why the universal indicator changes color from green to purple. (3 marks)

3. Give one piece of evidence which could be added to the table for all other reactants. (3 marks)

4. Write a word equation to describe the reaction between sodium and water. (3 marks)

5. Write a balanced chemical equation to describe the reaction between sodium and water. (3 marks)

6. The modern periodic table has been in use for over 150 years. It developed over the years with several scientists contributing to its development. (3 marks)

7. State one difference between the modern periodic table and Mendeleev’s table. (3 marks)

8. What is the periodicity of a group? (3 marks)

9. How can the periodic table be used to predict the properties of an element? (3 marks)

10. Explain why the position of an element in the periodic table indicates its reactivity. (3 marks)

Working scientifically: How theories change over time

After the discovery of the new element phosphorus in 1808, scientists began to think about the definition of an element. In the 1860s, Mendeleev produced a table similar to that below, to assign properties of elements. His table was based on their properties rather than their composition. It was not until after his death that the periodic nature of the table was recognized. The periodic nature of the table was recognized after the discovery of new elements such as chlorine and argon. These elements were assigned to elements with similar properties in the same group of the periodic table.

<table>
<thead>
<tr>
<th>Periodic table</th>
<th>Group 1 (alkali)</th>
<th>Group 2 (alkaline earth)</th>
<th>Group 3 (transition metals)</th>
<th>Group 4 (d-block)</th>
<th>Group 5 (p-block)</th>
<th>Group 6 (f-block)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Group 1</td>
<td>Lithium (Li)</td>
<td>Sodium (Na)</td>
<td>Potassium (K)</td>
<td>Rubidium (Rb)</td>
<td>cesium (Cs)</td>
<td>francium (Fr)</td>
</tr>
<tr>
<td>Group 2</td>
<td>Magnesium (Mg)</td>
<td>Calcium (Ca)</td>
<td>Strontium (Sr)</td>
<td>Barium (Ba)</td>
<td>Lanthane (La)</td>
<td>Cerium (Ce)</td>
</tr>
<tr>
<td>Group 3</td>
<td>Scandium (Sc)</td>
<td>Yttrium (Y)</td>
<td>Lutetium (Lu)</td>
<td>Hafnium (Hf)</td>
<td>Tantalum (Ta)</td>
<td>Niobium (Nb)</td>
</tr>
<tr>
<td>Group 4</td>
<td>Titanium (Ti)</td>
<td>Zirconium (Zr)</td>
<td>Hafnium (Hf)</td>
<td>Tantalum (Ta)</td>
<td>Niobium (Nb)</td>
<td>Molybdenum (Mo)</td>
</tr>
<tr>
<td>Group 5</td>
<td>Vanadium (V)</td>
<td>Chromium (Cr)</td>
<td>Manganese (Mn)</td>
<td>Iron (Fe)</td>
<td>Cobalt (Co)</td>
<td>Nickel (Ni)</td>
</tr>
<tr>
<td>Group 6</td>
<td>Copper (Cu)</td>
<td>Zinc (Zn)</td>
<td>Cadmium (Cd)</td>
<td>Mercury (Hg)</td>
<td>Thallium (Tl)</td>
<td>Lead (Pb)</td>
</tr>
<tr>
<td>Group 7</td>
<td>Silver (Ag)</td>
<td>Mercury (Hg)</td>
<td>Indium (In)</td>
<td>Tin (Sn)</td>
<td>Antimony (Sb)</td>
<td>Tellurium (Te)</td>
</tr>
<tr>
<td>Group 8</td>
<td>Cadmium (Cd)</td>
<td>Indium (In)</td>
<td>Thallium (Tl)</td>
<td>Tin (Sn)</td>
<td>Antimony (Sb)</td>
<td>Tellurium (Te)</td>
</tr>
<tr>
<td>Group 9</td>
<td>Mercury (Hg)</td>
<td>Indium (In)</td>
<td>Thallium (Tl)</td>
<td>Tin (Sn)</td>
<td>Antimony (Sb)</td>
<td>Tellurium (Te)</td>
</tr>
<tr>
<td>Group 10</td>
<td>Tellurium (Te)</td>
<td>Tellurium (Te)</td>
<td>Tellurium (Te)</td>
<td>Tellurium (Te)</td>
<td>Tellurium (Te)</td>
<td>Tellurium (Te)</td>
</tr>
</tbody>
</table>

In addition to many elements which form the basis of our modern periodic table, Mendeleev’s list also included “light” and “heavy” metals which at the time were believed to be earth substances. Mendeleev discovered that elements with similar properties could be classified into groups. The modern classification of these compounds as elements was due to a lack of technology as well as a lack of knowledge.
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 4.2.1 to 4.2.4 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 the Appendix.
Previously you could have learned:

- Substances exist in one of three states of matter: solid, liquid, or gas.
- In solids, the particles are vibrating and packed close together in a regular pattern; in liquids, the particles are close together but moving about randomly; in gases, the particles are far apart and moving about randomly.
- Substances change state at their melting and boiling points.
- How easily substances melt or boil depends on the strength of the forces or bonds between the particles.
- Different substances have different properties, e.g., some have high melting points but others have low melting points, some conduct electricity but many do not.

Test yourself on prior knowledge:
- What are the three states of matter?
- In general terms, what is the name of the temperature at which a substance changes from:
  - A solid to a liquid.
  - A liquid to a gas.
- Identify two substances in each case that:
  - Have a high melting point.
  - Have a low melting point.
  - Conduct electricity as a solid.
  - Do not conduct electricity as a solid.
- The particles in liquids and gaseous can move around, but those in a solid only vibrate and cannot move around. Explain why the particles in a solid cannot move around.

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 sulphate is made from copper (metal), sulphur (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 that have a grid-like lattice. A giant lattice contains a massive number of particles in a regular structure that continues in all directions throughout the substance (Figure 2.1).

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 2.2). 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 as they cannot conduct electricity. However, when melted, the ions can move and carry charge, so ionic substances will conduct electricity when melted. Many ions substances dissolve in water and...
The formula of ionic substances

The charge on ions
You can work out the charge on ions from the number of electrons that have left or gained. For example, all the elements in Group 1 have one electron in their outer shell and so lose one electron when they form ions (e.g., Na⁺, K⁺). Group 2 have two electrons in their outer shell and so lose two electrons when they form ions (e.g., Ca²⁺, Mg²⁺). Group 7 have seven electrons in their outer shell and so gain one electron when they form ions (e.g., Cl⁻, Br⁻).

These charges and those of other common ions are shown in the Tables 2.2 and 2.3.
Molecular substances

**What are molecular substances?**

Molecules are made of atoms bonded together by a special type of bond called a **covalent bond**. A molecule is a group of atoms held together by these covalent bonds. Not all substances are molecules; some are made of a single type of atom, such as hydrogen or oxygen, and these are called **atomic substances**. Examples of molecular substances include water, hydrogen, oxygen, and carbon dioxide.

- **KEY TERMS**
  - Molecular substance: Made of atoms joined by covalent bonds.
  - Covalent bond: A single bond that forms when two atoms share their outermost electrons.

**The structure of molecular substances**

A molecule is made up of its constituent atoms, which are held together by covalent bonds. The shape and properties of a molecule depend on the arrangement of these atoms and the nature of the bonds between them. For example, in a water molecule (H₂O), the two hydrogen atoms are bonded to the oxygen atom, forming a V-shape.

- **KEY TERMS**
  - Intermolecular forces: Weak forces between molecules.

**The properties of molecular substances**

- **Melting and boiling points**

  Molecules are not bonded to each other. The intermolecular forces (forces between molecules) are very weak and so are easy to overcome. This means that molecular substances have high 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.

**Electrical conductivity**

Molecules are electrically neutral, which means they do not conduct electricity.

**Show you can...**

- Test yourself: Explain why molecular substances have low melting and boiling points.

**Generally, the larger the molecule, the stronger the forces between the molecules and the higher the melting and boiling points.**

Molecules of glucose are quite large and it melts at 144°C. When molecular substances change state, the covalent bonds do not break. For example, water molecules are identical in size whether it is present, liquid, or in the gaseous state (Figure 2.10). No covalent bonds are broken when water changes state.

**The formula of molecular substances**

Molecular substances have two formulas, the empirical formula and the molecular formula. The molecular formula is the one that is normally used.
Polymers

There are many different types of polymer (plastics), including polyethene, PVC, Teflon, silicones, and polyester. Polymers contain very large molecules, often with hundreds or thousands of atoms. Within each molecule, the atoms are joined to each other by covalent bonds (Figure 2.13).

Giant covalent substances

- What are giant covalent substances?
  - There are 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 (Different forms of carbon section)
    - graphite, C (a form of carbon) – studied in detail later in the chapter (Different forms of carbon section)
    - silicon, Si, 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 and throughout the structure (Figure 2.13).
  - These substances are solid 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. In a giant covalent substance, there is one continuous network.
- The properties of giant covalent substances
  - Melting and boiling points: to order to melt a giant covalent substance, many covalent bonds have to be broken. Covalent bonds are very strong and it takes a lot of energy to break them. Therefore, giant covalent substances have very high melting and boiling points. For example, diamond melts at over 5900°C.
  - Electrical conductivity: most giant covalent substances do not conduct electricity because they do not contain any delocalized electrons. However, graphite does, as it does some delocalized electrons. Delocalized electrons are able to move through the substance.

Test yourself

1. Describe the structure of giant covalent substances.
2. Why do giant covalent substances have very high melting points?
3. Why do giant covalent substances, except graphite, not conduct electricity?
**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 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 2.14 and 2.15).

- **The properties of metallic substances**
  **Melting and boiling points**
  In metals, the metallic bonding is strong. This means that most metals 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 2.16).

**Thermal conductivity**
Metals are also good thermal conductors of heat. 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 2.17). This makes metals soft.

**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 mixing 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, steel is an alloy 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 different 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 2.18).

**Test yourself**
1. Why is copper a good conductor of electricity?
2. Why is aluminium better than copper?
3. Why do metals conduct electricity?
4. Why does a metal conduct electricity?
5. Why are metals malleable?
6. Why are metals not good conductors of electricity?
7. Why do metals conduct electricity?
8. Why are metals not good conductors of electricity?
9. Why are metals not good conductors of electricity?
10. Why are metals not good conductors of electricity?
Overview of types of bonding and structures

Types of bonding

There are three types of bonding that are summarized in the Table 2.8.

<table>
<thead>
<tr>
<th>Description</th>
<th>Ionic</th>
<th>Covalent</th>
<th>Metallic</th>
</tr>
</thead>
<tbody>
<tr>
<td>A bond formed by electrostatic forces between ions</td>
<td>A bond formed by sharing of electrons between atoms</td>
<td>A bond formed by delocalized electrons in a metal lattice</td>
<td></td>
</tr>
</tbody>
</table>

Which substances have this bonding?

- Ionic: NaCl, KCl
- Covalent: CO₂, H₂O
- Metallic: Al, Cu

Types of structure

There are five types of structure that are summarized in Table 2.9.

Table 2.9

<table>
<thead>
<tr>
<th>Substances</th>
<th>Molecules</th>
<th>Lattices</th>
<th>Metals</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acetic acid</td>
<td>MO (CO₂H)</td>
<td>Ionic</td>
<td>Non-conductive</td>
</tr>
<tr>
<td>Phosphorus</td>
<td>P₄</td>
<td>Covalent</td>
<td>Non-conductive</td>
</tr>
<tr>
<td>Sodium bicarbonate</td>
<td>NaHCO₃</td>
<td>Ionic</td>
<td>Conductive</td>
</tr>
<tr>
<td>Sodium chloride</td>
<td>NaCl</td>
<td>Ionic</td>
<td>Conductive</td>
</tr>
<tr>
<td>Sodium hydroxide</td>
<td>NaOH</td>
<td>Covalent</td>
<td>Conductive</td>
</tr>
</tbody>
</table>

Substances may be classified in terms of their physical properties. Use the table to answer the following questions:

- Which substance could be used as an ionic conductor? (Answer: NaCl)
- Which substance could be used as a covalent conductor? (Answer: CO₂)
- Which substance could not be diamagnetic? (Answer: NaCl)
- Which substance is a metal? (Answer: NaCl)
**States of matter**

The three states of matter are solid, liquid and gas (Figure 2.19). Substances change state at their melting and boiling points. A substance is:

- 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 the 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 one in Figure 2.20. It can help you understand how particles are arranged when substances are in each state.

---

**Figure 2.19** Changes of state.

**Figure 2.20**

**Test yourself**

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>O₂</td>
<td>573</td>
<td>973</td>
</tr>
<tr>
<td>N₂</td>
<td>–196</td>
<td>–210</td>
</tr>
<tr>
<td>H₂O</td>
<td>0</td>
<td>100</td>
</tr>
<tr>
<td>CO₂</td>
<td>–78</td>
<td>573</td>
</tr>
<tr>
<td>NaCl</td>
<td>802</td>
<td>1,413</td>
</tr>
</tbody>
</table>

- Which substance sublimes at room temperature?  
- Which substance sublimes at 100°C?  
- Which substance sublimes at 800°C?  
- Which substance is a liquid over its solid temperature range?

---

**Show you can...**

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>NaCl</td>
<td>802</td>
<td>1,413</td>
<td>ionic bond</td>
</tr>
<tr>
<td>CO₂</td>
<td>–78</td>
<td>573</td>
<td>covalent bond</td>
</tr>
<tr>
<td>H₂O</td>
<td>0</td>
<td>100</td>
<td>hydrogen bond</td>
</tr>
<tr>
<td>N₂</td>
<td>–196</td>
<td>–210</td>
<td>closed packed structure</td>
</tr>
<tr>
<td>O₂</td>
<td>573</td>
<td>973</td>
<td>open packed structure</td>
</tr>
<tr>
<td>NaCl</td>
<td>802</td>
<td>1,413</td>
<td>ionic bond</td>
</tr>
</tbody>
</table>

- Which element is a solid at room temperature?  
- Which compound is a gas at room temperature?  
- Which element sublimes when cooled to room temperature?  
- Which compound will freeze first on cooling from room temperature to a very low temperature?
**Nanoscience**

**What is nanoscience?**

Nanoscience is the study of nanomaterials. Nanoparticles are structures that are between 1 nm and 100 nm in size. Nanoparticles typically contain a few hundred atoms. Table 2.10 compares the size of nanoparticles to other particles.

<table>
<thead>
<tr>
<th>Size (nm)</th>
<th>Nanoparticle</th>
<th>Other Particles</th>
</tr>
</thead>
<tbody>
<tr>
<td>1-100</td>
<td>True</td>
<td>False</td>
</tr>
</tbody>
</table>

**Nanomaterials**

In nanomaterials, there is a huge number of gold atoms in a single particle. Only a tiny fraction of the gold atoms are on the surface of the particle.

Nanomaterials of gold can be made and contain a few hundred gold atoms (Figure 2.2). In the nanoparticles, a much higher fraction of the atoms are on the surface.

Bulk gold metal is very unreactive. However, gold nanoparticles have different properties and can move sides. For example, gold nanoparticles can be used to catalyze some chemical reactions.

**Why do nanoparticles behave differently to the bulk material?**

In a bulk material, such as a piece of gold, there is one large structure where only a tiny fraction of the atoms is on the surface. However, in a nanoparticle, which is much smaller, a much higher fraction of the atoms are on the surface. This explains why nanoparticles have different properties to the bulk material.

A way to think about this is to look at the surface area to volume ratio of the structure. In Table 2.11, we are comparing a large cube with one that has sides that are ten times shorter in length. As the sides of a cube get shorter by a factor of 10, the surface area to volume ratio increases by a factor of 10. Again this shows that in a smaller structure such as a nanoparticle, a higher fraction of the material is at the surface than in the larger bulk material.

**Uses of nanoparticles**

There are many uses for nanoparticles and finding new applications of nanoparticles is a very active and important area of scientific research. Some applications of nanoparticles at present are discussed here.

**In fuel cells**

Fuel cells are very important and believed to be vital in the years to come as an energy source (Figure 2.2). They use the electrochemical reaction between hydrogen and oxygen to release electrical energy. Most fuel cells use platinum metal as a catalyst for the reaction, but platinum is very expensive. Nanoparticles of platinum are being used in two ways to lower the cost of the fuel cells. Firstly, platinum nanoparticles can be used as the catalyst, meaning that less platinum is needed. Secondly, nanoparticles of alternative metals are being developed to replace the platinum.

**Delivery of drugs**

Nanoparticles can be used to deliver drugs to specific cells in the body. This will reduce the amount of the drug needed and reduce any side effects of the drug. Gold nanoparticles are one example of nanoparticles that are being used in this way.

**In sun creams**

Nanoparticles of titanium dioxide (TiO2) or zinc oxide (ZnO) are used to absorb harmful UV radiation in sun creams. They give better skin coverage and better protection from UV radiation than normal sun creams, but they are also clear and colorless in the traditional sun creams (Figure 2.22).

**Synthetic skin**

Nanoparticles are used to treat patients with severe burns. Nanoparticles (e.g., carbon nanotubes) are being used to create better synthetic skin.
that is stronger but more flexible. Research is even taking place to use nanoparticles to create synthetic skin that senses touch and heat.

**Cosmetics**

Some examples of the use of nanoparticles in cosmetics include:
- In face creams in emulsions that contain vitamins
- In formulations to kill bacteria
- In foundations to diffuse light to partially disguise wrinkles.

**Clothing**

Some clothes contain silver nanoparticles. Examples include socks, sports clothing, underwear, and zippers. The silver nanoparticles kill bacteria preventing the build-up of unpleasant odors.

**Deodorants**

Some deodorants also contain silver nanoparticles which kill bacteria preventing the build-up of unpleasant odors.

**Electronics**

Nanoparticles are being used to improve electronic components (Figure 2.24). For example, some circuit boards are now printed using nanoparticles and some chips now contain nanoparticles. The use of nanoparticles is allowing smaller components to be made.

- **The safety of nanoparticles**

There are some concerns about the use of nanoparticles. It might have been assumed that if a bulk material is safe to use then nanoparticles of that material would be also. However, if the nanoparticles behave differently to the bulk material, then it is reasonable to assume nanoparticles may well have harmful effects on humans or the environment that the bulk material does not. For example, bulk gold is safe but gold nanoparticles, which have many uses, may well not be safe. One way in which nanoparticles could be harmful is that some may be able to penetrate cell membranes to enter cells whereas the bulk material may not. There is some evidence that some nanoparticles may cause problems even though the bulk material does not.

**Show you can...**

- State 4 reasons why it is a good idea that they contain nanoparticles. 
- Explain what is meant by the term “nanotechnology.”
- Explain why it is important to use nanoparticles in the manufacture of
  - electronics and
  - cosmetics.
- Explain why it is possible to use nanoparticles in the manufacture of
  - drugs.
- Explain why it is possible to use nanoparticles in the manufacture of
  - food.
- Explain why it is possible to use nanoparticles in the manufacture of
  - clothing.
- Explain why it is possible to use nanoparticles in the manufacture of
  - deodorants.
- Explain why it is possible to use nanoparticles in the manufacture of
  - electronics.
- Explain why it is possible to use nanoparticles in the manufacture of
  - cosmetics.
- Explain why it is possible to use nanoparticles in the manufacture of
  - drugs.
- Explain why it is possible to use nanoparticles in the manufacture of
  - food.
- Explain why it is possible to use nanoparticles in the manufacture of
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- Explain why it is possible to use nanoparticles in the manufacture of
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  - electronics.
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>
<thead>
<tr>
<th>Table 2.12 Comparison of the physical properties of diamond and graphite</th>
</tr>
</thead>
<tbody>
<tr>
<td>Property</td>
</tr>
<tr>
<td>Melting point</td>
</tr>
<tr>
<td>Hardness</td>
</tr>
<tr>
<td>Conductivity</td>
</tr>
<tr>
<td>Malleability</td>
</tr>
<tr>
<td>Brittle</td>
</tr>
<tr>
<td>Oxidation</td>
</tr>
</tbody>
</table>

**Graphene**

Graphene is a new substance. It is a single layer of graphite (Figure 2.29). 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 carbon atom thick, but it is incredibly strong due to its giant covalent structure. It is also a superb conductor of electricity. 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 fibre).

**Fullerenes**

The molecules C₆₀ were identified in the 1980s as another form of carbon. The molecule has a shape that resembles a football (Figure 2.29). It was named buckminsterfullerene after the architect Richard Buckminster Fuller who built domes that had similar structures. C₆₀ is often referred to as a buckyball and was the first fullerene produced. Scientists at the University of Sussex won a Nobel Prize in 2005 for their work on fullerenes.

**Carbon nanotubes**

Carbon nanotubes are cylindrical fullerenes, sometimes called buckytubes (Figure 2.29). 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. 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 conductivity – 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 rackets (Figure 2.30) and golf clubs.

**Show you can…**

Copy and complete the table to give information about the different substances and use of carbon.
Chapter review questions

1. What structure types are: a) ionic, b) metallic, c) molecular, d) 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 freezing point.
   b. Explain why carbon dioxide does not conduct electricity.
   c. Why is it a gas, not a liquid at room temperature?

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. Explain why steels are harder than pure iron.

4. Potassium fluoride is an ionic compound containing potassium cations and fluoride anions.
   a. Give the electron structure of potassium (K⁺) ions.
   b. Give the electron 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 melted or dissolved.

5. Decide whether each of the following substances has an ionic, molecular, giant covalent or metallic structure:
   a. H₂O
   b. F₂
   c. MgO
   d. CCl₄
   e. NaCl
   f. MgCl₂
   g. NaClO₃
   h. CaCl₂
   i. Na₂O
   j. NaClO
   k. Na₂CO₃
   l. NaOH

6. Given the following table, determine the melting point and conductivity:

<table>
<thead>
<tr>
<th>Substance</th>
<th>Melting point (°C)</th>
<th>Conductivity</th>
</tr>
</thead>
<tbody>
<tr>
<td>X</td>
<td>65°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>Y</td>
<td>80°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>Z</td>
<td>120°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>A</td>
<td>20°C</td>
<td>Conducts</td>
</tr>
<tr>
<td>B</td>
<td>40°C</td>
<td>Conducts</td>
</tr>
<tr>
<td>C</td>
<td>150°C</td>
<td>Conducts</td>
</tr>
<tr>
<td>D</td>
<td>180°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>E</td>
<td>190°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>F</td>
<td>200°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>G</td>
<td>210°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>H</td>
<td>220°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>I</td>
<td>230°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>J</td>
<td>240°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>K</td>
<td>250°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>L</td>
<td>260°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>M</td>
<td>270°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>N</td>
<td>280°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>O</td>
<td>290°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>P</td>
<td>300°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>Q</td>
<td>310°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>R</td>
<td>320°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>S</td>
<td>330°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>T</td>
<td>340°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>U</td>
<td>350°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>V</td>
<td>360°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>W</td>
<td>370°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>X</td>
<td>380°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>Y</td>
<td>390°C</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>Z</td>
<td>400°C</td>
<td>Does not conduct</td>
</tr>
</tbody>
</table>

7. Draw a sketch and a dot-cross diagram for each of the following molecules:
   a. CH₄
   b. NO₂
   c. CO₂
   d. HF
   e. H₂O

8. Calcium reacts with chlorine to form the ionic compound calcium chloride. Draw a sketch to show the electron structure in calcium atoms, chlorine atoms, calcium chloride, and the formation of the bond.

9. Work out the formulas of the following ionic compounds. The charge of some ions is given (F⁺ = +1, Cl⁻ = -1, Na⁺ = +1, K⁺ = +1, etc.).
   a. potassium oxide
   b. magnesium fluoride
   c. lithium sulfide
   d. nickel(II) chloride
   e. copper(II) nitrate

10. Polypropene is a polymer made of molecules which typically melt at 130°C.
    a. Describe the bonding within the polymer molecules.
    b. Explain why polypropene has a relatively high melting point for a molecular substance.

11. Gold nanoparticles have different properties to bulk gold.
    a. What are nanoparticles?
    b. Explain why nanoparticles often have different properties from the bulk material.
    c. Why are some people concerned about the use of nanoparticles?
    d. Calculate the number of copper atoms in 0.1 mol of copper at room temperature.

12. Ethylene is a molecule with the molecular formula C₂H₄.
    a. Explain what this formula tells us about ethylene.
    b. Show a dot-cross diagram to show the covalent bonds in a molecule of ethylene.
    c. Show a box-cross diagram to show the outer shell electrons in a molecule of ethylene.

13. The diagram shows part of a carbon molecule. They are used to reinforce the materials in the making of carbon fibre composites to have low and high tensile strength.
    a. How many resonance bonds does each carbon atom make in a carbon molecule?
    b. Explain briefly by considering your answer to c) why molecules can conduct electricity.
    c. Explain briefly why carbon molecules have high tensile strength.
Practice questions

1. The elements A, B and C are 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 between them? [1 mark]
   - A: Mg, covalent
   - B: Pb, ionic
   - C: Pb, metallic
   - D: Mg, metallic

2. Which one of the following does not have a giant covalent structure? [1 mark]
   - A: diamond
   - B: graphite
   - C: boron
   - D: sodium chloride

3. A dot and cross diagram is shown here.

   [Diagram of A dot and cross diagram]

   a) Name and write the formula for this compound. [2 marks]
   b) On the diagram above, use an arrow to label:
      1. A covalent bond.
      3. Using a line to represent a single covalent bond, redraw the diagram shown above. [3 marks]
   c) What is meant by the term single covalent bond? [2 marks]
   d) Non-covalence is the study of non-covalent bonds. [1 mark]

4. Non-covalence is the study of non-covalent bonds. [1 mark]
   a) What is the name of particles studied in nanotechnology? [1 mark]
   b) Nanoparticles can be added to other materials. Adding nanoparticles changes the properties of these materials. Describe two examples of products, other than those already known, that have nanoparticles added to them. Explain how adding nanoparticles changes the properties of these products, and suggest why this is useful. [6 marks]

5. The following diagram shows some changes between the states of matter.

   - A: solid
   - B: liquid
   - C: gas

   a) What is the name for each of the changes labeled A, B, and C? [3 marks]
   b) Cold water conducts electricity. Explain why. [2 marks]

6. 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]

7. Explain why sodium chloride solution will conduct electricity, but sodium chloride solid will not. [3 marks]

8. The table gives some of the properties of the Period 3 element magnesium and one of its compounds, magnesium chloride.

<table>
<thead>
<tr>
<th>Property</th>
<th>Magnesium</th>
<th>Magnesium chloride</th>
</tr>
</thead>
<tbody>
<tr>
<td>Melting point (°C)</td>
<td>650</td>
<td>710</td>
</tr>
<tr>
<td>Density (g/cm³)</td>
<td>1.74</td>
<td>2.71</td>
</tr>
<tr>
<td>Electrical conductivity</td>
<td>does not conduct</td>
<td>conducts</td>
</tr>
</tbody>
</table>

   a) Which two of the diagrams represent different structures of the element carbon? [3 marks]
   b) The substances are carbon dioxide, diamond, graphite or chlorine dioxide. Name each substance, A to D. [4 marks]
   c) Name the type of bonding which connects between the atoms in all of the substances A to D. [3 marks]
   d) Name the type of structure for each substance A to D. [4 marks]

9. The photographs show two cases of graphite. It is used in pencils and for electrodes in the electrolysis of sodium carbonate solution.

   a) Which reference to the structure of graphite explains why it is used in pencils? [3 marks]
   b) Explain why graphite is a good conductor of electricity. [2 marks]

10. a) Name the type of structure for calcium chloride. [1 mark]
    b) Use a dot and cross diagram to show how atoms of calcium combine with atoms of chlorine to form calcium chloride. [3 marks]
    c) Name the type of bonding in calcium chloride. [1 mark]

11. a) Name the type of structure for calcium chloride. [1 mark]
    b) Use a dot and cross diagram to show how atoms of calcium combine with atoms of chlorine to form calcium chloride. [3 marks]
    c) Name the type of bonding in calcium chloride. [1 mark]

12. a) Name the type of structure for calcium chloride. [1 mark]
    b) Use a dot and cross diagram to show how atoms of calcium combine with atoms of chlorine to form calcium chloride. [3 marks]
    c) Name the type of bonding in calcium chloride. [1 mark]

13. The properties of compounds depend very clearly on their bonding. Redraw the following table with only the correct words to show some of the properties of calcium chloride and sodium carbonate. [3 marks]

<table>
<thead>
<tr>
<th>Compound</th>
<th>Solubility in water</th>
<th>Electrical conductivity</th>
</tr>
</thead>
<tbody>
<tr>
<td>Calcium chloride</td>
<td>soluble</td>
<td>does not conduct</td>
</tr>
<tr>
<td>Sodium carbonate</td>
<td>soluble</td>
<td>does not conduct</td>
</tr>
</tbody>
</table>

14. a) Name the type of structure for calcium chloride. [1 mark]
    b) Use a dot and cross diagram to show how atoms of calcium combine with atoms of chlorine to form calcium chloride. [3 marks]
    c) Name the type of bonding in calcium chloride. [1 mark]

15. a) Name the type of structure for calcium chloride. [1 mark]
    b) Use a dot and cross diagram to show how atoms of calcium combine with atoms of chlorine to form calcium chloride. [3 marks]
    c) Name the type of bonding in calcium chloride. [1 mark]

16. a) Name the type of structure for calcium chloride. [1 mark]
    b) Use a dot and cross diagram to show how atoms of calcium combine with atoms of chlorine to form calcium chloride. [3 marks]
    c) Name the type of bonding in calcium chloride. [1 mark]

17. a) Name the type of structure for calcium chloride. [1 mark]
    b) Use a dot and cross diagram to show how atoms of calcium combine with atoms of chlorine to form calcium chloride. [3 marks]
    c) Name the type of bonding in calcium chloride. [1 mark]

18. a) Name the type of structure for calcium chloride. [1 mark]
    b) Use a dot and cross diagram to show how atoms of calcium combine with atoms of chlorine to form calcium chloride. [3 marks]
    c) Name the type of bonding in calcium chloride. [1 mark]

19. a) Name the type of structure for calcium chloride. [1 mark]
    b) Use a dot and cross diagram to show how atoms of calcium combine with atoms of chlorine to form calcium chloride. [3 marks]
    c) Name the type of bonding in calcium chloride. [1 mark]

20. a) Name the type of structure for calcium chloride. [1 mark]
    b) Use a dot and cross diagram to show how atoms of calcium combine with atoms of chlorine to form calcium chloride. [3 marks]
    c) Name the type of bonding in calcium chloride. [1 mark]
**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 manipulated. It is easier to say that a speed of dust has a mass of 1.2 x 10^16 grams then to say it has a mass of 1,200,000,000,000,000 grams, even though they mean the same thing.

*Standard form must always look like this:*

\[ a \times 10^n \]

- **a** must always be between 1 and 10
- \( n \) is the number of places the decimal point moves
  - \( n \) is positive for numbers greater than 1
  - \( n \) is negative for numbers less than 1
  - The value of \( a \) will be a number

**Example**
Write 0.000 000 000 000 001 in standard form.

**Answer**

- \( a \) is the first non-zero digit with a decimal place after the first number and then \( n \) = 10 after it
- \( 4 \times 10^{-10} \)

**Example**
Write 6.0 x 10^5 in standard form.

**Answer**

- \( a \) is the first non-zero digit with a decimal place after the first number and then \( n \) = 10 after it
- \( 6 \times 10^5 \)

**SI Units**
The International System of Units (SI) is a system of units of measurements that is widely used around the world. Just like you would use a metric unit in your study of chemistry, it uses several base units of measure. For example, systems, grams and seconds, when a numerical is very small or large, the units may be modified by using a prefix. Some prefixes are shown in the table.

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>kilo</td>
<td>k</td>
<td>1,000</td>
</tr>
<tr>
<td>mega</td>
<td>M</td>
<td>1,000,000</td>
</tr>
<tr>
<td>giga</td>
<td>G</td>
<td>1,000,000,000</td>
</tr>
<tr>
<td>nano</td>
<td>n</td>
<td>0.000 000 000 1</td>
</tr>
<tr>
<td>micro</td>
<td>µ</td>
<td>0.000 000 1</td>
</tr>
<tr>
<td>pico</td>
<td>p</td>
<td>0.000 000 000 001</td>
</tr>
</tbody>
</table>

**Questions**

1. Write the numbers below in standard form.
   - 0.000 000 000 000 001
   - 1,000 000
   - 0.000 000 000 000 001
   - 1,000 000
   - 0.000 000 000 000 001
   - 1,000 000

2. Write the numbers below in scientific notation.
   - 6 x 10^-7
   - 3 x 10^-8
   - 4 x 10^-9
   - 1 x 10^-5
   - 2 x 10^-6
   - 3 x 10^-7

3. Each prefix has a symbol that helps you identify the units represented by that prefix. Write the symbol that represents each unit:
   - 1 km
   - 1 mg
   - 1 mm
   - 1 m

4. Each prefix has a symbol that helps you identify the units represented by that prefix. Write the symbol that represents each unit:
   - 1 km
   - 1 mg
   - 1 mm
   - 1 m

5. Write the following quantities in SI units with the appropriate prefixes.
   - 1,234 m
   - 0.000 000 001 m
   - 1,234 cm
   - 0.000 000 001 cm
   - 1,234 mm
   - 0.000 000 001 mm

6. The sizes of objects vary in size. What is the size of a virus?
   - The sizes of objects vary in size. What is the size of a virus? If the size of a virus is less than 1 mm, what is its size in nanometers and meters?

7. The size of a virus is less than 1 mm, what is its size in nanometers and meters?

8. A single carbon atom has a diameter of 1.5 x 10^-7 m. Write this using a prefix.
This chapter covers specification points 4.3.1 to 4.3.3 and is called Quantitative chemistry. It also covers how to carry out a titration from 4.4.2.

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, including titrations, 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 and how to carry out titrations.

Relative mass and moles

- Relative atomic mass
  Individual atoms have a tiny mass. For example, an atom of H has a mass of about $2 \times 10^{-23}$ g (that is 0.000000000000000000002 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 H atom is defined as being exactly 1.6. On this scale, an atom of Mg has a relative mass of 24 and is twice as heavy as a H atom. Similarly, an atom of N has a relative mass of 14 and is 8 times heavier than a H 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 $^{35}$Cl with relative mass 35 and the remaining 25% are $^{37}$Cl atoms with relative mass 37. The average relative mass of chlorine atoms is 35.5 as there are more chlorine atoms with relative mass 35 than 37.
  The relative atomic mass (A) of an element is the average mass of atoms of that element taking into account the mass and amount of each isotope it contains on a scale where the mass of a H atom is 1.

- Relative formula mass
  The relative formula mass (M) of a substance is the sum of the relative atomic masses of all the atoms shown in the formula. It is often just called formula mass.
**Key Terms**

Relative formula mass: The sum of the relative atomic masses of all the atoms in one molecule of the compound (when referred to as formula mass).

**Table 3.1**

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>A. number</th>
<th>Sum M</th>
<th>R. from Hg</th>
<th>R. from metal</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>H₂O</td>
<td>18.018</td>
<td>1818</td>
<td>0.02</td>
<td>10</td>
</tr>
<tr>
<td>Copper</td>
<td>Cu</td>
<td>63.5</td>
<td>655</td>
<td>0.01</td>
<td>10</td>
</tr>
<tr>
<td>Sodium chloride</td>
<td>NaCl</td>
<td>58.5</td>
<td>595</td>
<td>0.01</td>
<td>10</td>
</tr>
<tr>
<td>Sulfuric acid</td>
<td>H₂SO₄</td>
<td>98.075</td>
<td>981</td>
<td>0.02</td>
<td>10</td>
</tr>
<tr>
<td>Magnesium oxide</td>
<td>MgO</td>
<td>40.31</td>
<td>413</td>
<td>0.01</td>
<td>10</td>
</tr>
<tr>
<td>Ammonium nitrate</td>
<td>NH₄NO₃</td>
<td>80.05</td>
<td>810</td>
<td>0.01</td>
<td>10</td>
</tr>
</tbody>
</table>

**Show you can...**

Many compounds can exist in different physical states (solid, liquid, gas). The simplest way to describe these states is in terms of their formula mass.

**How many moles?**

If one mole of H₂O atoms has a mass of 18.02 g, then it follows that the mass of two moles of H₂O atoms will be 36.04 g. There is a simple equation linking the mass of a substance to the number of moles (Figure 3.2).

When using this equation, the mass must be in grams. Table 3.3 shows conversion factors if the moles are not given in grams.

**Table 3.2**

<table>
<thead>
<tr>
<th>Compound</th>
<th>Formula</th>
<th>Relative formula mass (M)</th>
<th>Molar mass (Mₐ)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H₂O</td>
<td>18.018</td>
<td>18.02 g/ mol</td>
<td></td>
</tr>
<tr>
<td>NaCl</td>
<td>58.5</td>
<td>58.50 g/ mol</td>
<td></td>
</tr>
<tr>
<td>H₂SO₄</td>
<td>98.075</td>
<td>98.08 g/ mol</td>
<td></td>
</tr>
<tr>
<td>MgO</td>
<td>40.31</td>
<td>40.31 g/ mol</td>
<td></td>
</tr>
<tr>
<td>NH₄NO₃</td>
<td>80.05</td>
<td>80.06 g/ mol</td>
<td></td>
</tr>
</tbody>
</table>

**Table 3.3**

<table>
<thead>
<tr>
<th>Compound</th>
<th>Molar mass (Mₐ)</th>
<th>Conversion factors</th>
</tr>
</thead>
<tbody>
<tr>
<td>H₂O</td>
<td>18.02 g/ mol</td>
<td>1 g/ 0.054 mol</td>
</tr>
<tr>
<td>NaCl</td>
<td>58.50 g/ mol</td>
<td>1 g/ 0.017 mol</td>
</tr>
<tr>
<td>H₂SO₄</td>
<td>98.08 g/ mol</td>
<td>1 g/ 0.010 mol</td>
</tr>
<tr>
<td>MgO</td>
<td>40.31 g/ mol</td>
<td>1 g/ 0.025 mol</td>
</tr>
<tr>
<td>NH₄NO₃</td>
<td>80.06 g/ mol</td>
<td>1 g/ 0.013 mol</td>
</tr>
</tbody>
</table>

**Test yourself**

1. Calculate the molar mass of the following substances:
   - H₂O
   - NaCl
   - H₂SO₄
   - MgO
   - NH₄NO₃

2. A 5.00 g sample of H₂O contains how many moles of H₂O?

3. How many grams of H₂SO₄ are required to prepare 1.00 mol of H₂SO₄?

4. A sample of MgO contains 0.050 mol of MgO. How many grams of MgO are in the sample?

5. A 2.50 g sample of NH₄NO₃ contains how many moles of NH₄NO₃?

Table 3.4 gives some examples using the equation that links mass and moles. In calculations, the units of moles is usually abbreviated to mol.
Test yourself

Use the figure 2.3 to write down the atomic mass in the table at the back of this book. How many atoms of the following compounds are there in a molecule?

1. 2 H atoms
2. 1 N atom
3. 3 O atoms

Fill in the blanks below.

- 2 H atoms
- 1 N atom
- 3 O atoms

Significant figures

We often quote answers to calculations to a certain number of significant figures. In chemistry, we usually quote values to 2, 3, or 4 significant figures (SF), but it can be more or less than this, Table 3.5 shows some numbers given to 2, 3, and 4 significant figures.

We count 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 21.3333333°C. However, it is impossible for us to say that the temperature rise is exactly 21.3333333°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.

Test yourself

1. How many molecules of oxygen are in 7.6g of oxygen?
2. How many molecules of water are in 5.4g of water?
3. How many molecules of carbon dioxide are in 16.2g of carbon dioxide?
4. How many molecules of nitric oxide are in 14.8g of nitric oxide?

Conservation of mass

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 3.1, one molecule of nitrogen (N2) reacts with three molecules of hydrogen (H2) to make two molecules of ammonia (NH3).

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 of the reactants in the quantities shown in the equation will add up to the total of the relative formula masses of all of the products in the quantities shown in the equation (Figure 3.4).

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 do so 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 "lost" 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 3.5, some magnesium has been reacted with a carbon dioxide, which results in the formation of magnesium oxide and carbon dioxide.
Thermal decomposition reactions

A thermal decomposition reaction is one where heat causes a substance to break down into simpler substances.

TIP
Metal carbonates decompose into a metal oxide and carbon dioxide when heated.

Reaction

Before reaction

After reaction

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 lot of oxygen. During heating the crucible was tilted from time to time.

The following results were obtained:

Mass of crucible = 13.36 g
Mass of crucible + titanium metal = 17.28 g
Mass of crucible + titanium oxide = 15.02 g

Questions

1. Use the results to calculate:
   - the mass of titanium oxide formed in this experiment.
   - the mass of oxygen used in this experiment.
   - the equation for the reaction in terms of titanium, oxygen and titanium oxide.
2. Suggest why it was necessary to tilt the crucible during heating.

Test yourself

You can find a further set of exam questions on the periodic table at the back of this book to help you answer these questions.

10. Hydrogen reacts with oxygen to make water as shown in this equation:

   \[ \text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O} \]

   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 molar masses of all the reactants equals the sum of the molar masses of all the products.
   c) A piece of copper was heated in air. After a few minutes it was weighed, and found to be heavier.
   d) Explain why the copper gets heavier.
   e) What is the law of conservation of mass?
   f) Explain why this reaction does not break the law of conservation of mass.
   g) When 1.02 g of solid nickel oxide is heated for several minutes, only 0.72 g of nickel metal is obtained. Explain why the mass decreases and what happens to the remaining mass.

Molar ratios in equations

Chemical equations can be interpreted in terms of ratios. For example, in the equation in Figure 3.4 one mole of nitrogen (N₂) molecules reacts with three moles of hydrogen (H₂) molecules to form two moles of ammonia (NH₃) molecules.

1. For each of the following reactions:
   - 2N₂ + 3H₂ \rightarrow 2NH₃
   - C + O₂ \rightarrow CO₂
   - 4Fe + 3O₂ \rightarrow 2Fe₂O₃
   - 2Al + 3Cl₂ \rightarrow 2AlCl₃

   a) Write a balanced chemical equation for each reaction.
   b) Write the molar ratios involved for each reaction.
   c) Calculate the mass of reactants and products.
   d) Show that the molar ratios are consistent.

Figure 3.4 Balanced chemical equations and molar ratios in which substances react.
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 mass and relative formula masses.

Test yourself

1. Hydrogen reacts with oxygen to make water as shown in the equation below. What mass of oxygen is needed to react with 14 g of hydrogen?

$$\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}$$

2. Calcium carbonate decomposes when heated as shown in the equation below. What mass of calcium carbonate is needed to produce 0.20 g of calcium oxide?

$$\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$$

Using moles

1. How many moles of hydrogen react with 14 g of oxygen to make water?

2. How many moles of calcium carbonate are needed to produce 0.20 g of calcium oxide?

Using relative formula mass

1. Calculate the mass of magnesium oxide that can be produced from 14 g of magnesium.

2. What mass of calcium is needed to produce 0.20 g of calcium oxide?
Deducing the balancing numbers in an equation from reacting masses

The balancing numbers in a chemical equation can be calculated by calculating the masses of the substances in the reaction. In order to do this:

1. Calculate the mass of each substance (using molar mass, etc.).
2. Put this simplest whole number ratio of those mass values by dividing all the mass values by the smallest mass value.
3. If this does not give a whole number ratio, multiply up by a factor of 2 (if there is a value ending in approximately 0.2, or 0.5, or 0.7, or 0.65, etc., etc.).
4. Calculate the mass of sodium carbonate which is produced.

Example

1 g of magnesium reacts with 1 g of oxygen to make 2 g of magnesium oxide. Use this information to calculate the equation for this reaction.

Answer

<table>
<thead>
<tr>
<th>Substance</th>
<th>Mass (g)</th>
<th>Mass (g)</th>
<th>Mass (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Magnesium (Mg)</td>
<td>1</td>
<td>1</td>
<td>2</td>
</tr>
<tr>
<td>Oxygen (O₂)</td>
<td>32</td>
<td>32</td>
<td>32</td>
</tr>
<tr>
<td>Magnesium Oxide (MgO)</td>
<td>42</td>
<td>42</td>
<td>42</td>
</tr>
</tbody>
</table>

Therefore the reacting ratio is 1:1:2, and so the balanced equation is:

$\text{Mg} + \text{O}_2 \rightarrow \text{MgO}$

Example

8 g of aluminium reacts with 3.2 g of chlorine to make 22.4 g of aluminium chloride. Use this information to deduce the equation for this reaction.

Answer

<table>
<thead>
<tr>
<th>Substance</th>
<th>Mass (g)</th>
<th>Mass (g)</th>
<th>Mass (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aluminium (Al)</td>
<td>8</td>
<td>8</td>
<td>8</td>
</tr>
<tr>
<td>Chlorine (Cl₂)</td>
<td>3.2</td>
<td>6.4</td>
<td>6.4</td>
</tr>
<tr>
<td>Aluminium Chloride (AlCl₃)</td>
<td>32.4</td>
<td>32.4</td>
<td>32.4</td>
</tr>
</tbody>
</table>

Therefore the reacting ratio is 2:1:3, and so the balanced equation is:

$2\text{Al} + 3\text{Cl}_2 \rightarrow 2\text{AlCl}_3$
**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 consumed. This is often done if one of the reactants is readily available but the other one is expensive or in limited supply. An example of this is using excess air and oxygen to cause a complete combustion reaction. 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. A solution of magnesium (Mg) is reacted with 7 moles of sulfuric acid (H₂SO₄). One of the reactants is in excess. This is done to produce 2 moles of magnesium sulfate (MgSO₄) and 1 mole of H₂. The reaction is balanced as shown.

Mg + H₂SO₄ → MgSO₄ + H₂

**Reaction ratio from the equation**

<table>
<thead>
<tr>
<th>Reactant</th>
<th>Coefficient</th>
<th>Molarity</th>
<th>Moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mg</td>
<td>1</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>H₂SO₄</td>
<td>1</td>
<td>7</td>
<td>7</td>
</tr>
<tr>
<td>MgSO₄</td>
<td>1</td>
<td>1</td>
<td>2</td>
</tr>
<tr>
<td>H₂</td>
<td>1</td>
<td>1</td>
<td>1</td>
</tr>
</tbody>
</table>

**Reagent that takes place**

<table>
<thead>
<tr>
<th>Reactant</th>
<th>Moles of Reactant</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mg</td>
<td>1</td>
</tr>
<tr>
<td>H₂SO₄</td>
<td>7</td>
</tr>
<tr>
<td>MgSO₄</td>
<td>2</td>
</tr>
<tr>
<td>H₂</td>
<td>1</td>
</tr>
</tbody>
</table>

**Answer**

Magnesium reacts with sulfuric acid to produce 2 moles of magnesium sulfate (MgSO₄) and 1 mole of H₂. The reaction is balanced as shown.

Mg + H₂SO₄ → MgSO₄ + H₂

**Example**

Iron reacts with carbon monoxide as shown below. A solution of iron oxide (Fe₂O₃) is reacted with 5 moles of carbon monoxide (CO). The reaction is balanced as shown.

Fe₂O₃ + 3CO → 2Fe + 3CO₂

**Reaction ratio from the equation**

<table>
<thead>
<tr>
<th>Reactant</th>
<th>Coefficient</th>
<th>Molarity</th>
<th>Moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fe₂O₃</td>
<td>1</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>CO</td>
<td>3</td>
<td>5</td>
<td>15</td>
</tr>
<tr>
<td>2Fe</td>
<td>2</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>3CO₂</td>
<td>3</td>
<td>0</td>
<td>0</td>
</tr>
</tbody>
</table>

**Reagent that takes place**

<table>
<thead>
<tr>
<th>Reactant</th>
<th>Moles of Reactant</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fe₂O₃</td>
<td>1</td>
</tr>
<tr>
<td>CO</td>
<td>15</td>
</tr>
<tr>
<td>2Fe</td>
<td>0</td>
</tr>
<tr>
<td>3CO₂</td>
<td>0</td>
</tr>
</tbody>
</table>

**Answer**

Iron reacts with carbon monoxide to produce 2 moles of iron (Fe) and 3 moles of carbon dioxide (CO₂). The reaction is balanced as shown.

Fe₂O₃ + 3CO → 2Fe + 3CO₂

**Example**

A solution of tungsten (W) is reacted with 2 moles of hydrogen (H₂) to produce 2 moles of tungsten (W) and 4 moles of water (H₂O). The reaction is balanced as shown.

2W + 6H₂ → 4H₂O + 2W

**Reaction ratio from the equation**

<table>
<thead>
<tr>
<th>Reactant</th>
<th>Coefficient</th>
<th>Molarity</th>
<th>Moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>W</td>
<td>2</td>
<td>2</td>
<td>2</td>
</tr>
<tr>
<td>H₂</td>
<td>6</td>
<td>4</td>
<td>4</td>
</tr>
<tr>
<td>H₂O</td>
<td>4</td>
<td>0</td>
<td>0</td>
</tr>
</tbody>
</table>

**Reagent that takes place**

<table>
<thead>
<tr>
<th>Reactant</th>
<th>Moles of Reactant</th>
</tr>
</thead>
<tbody>
<tr>
<td>W</td>
<td>2</td>
</tr>
<tr>
<td>H₂</td>
<td>6</td>
</tr>
<tr>
<td>H₂O</td>
<td>0</td>
</tr>
</tbody>
</table>

**Answer**

A solution of tungsten (W) is reacted with 2 moles of hydrogen (H₂) to produce 2 moles of tungsten (W) and 4 moles of water (H₂O). The reaction is balanced as shown.

2W + 6H₂ → 4H₂O + 2W

**Figure 2.11 Tungsten**
Yield and atom economy

Percentage yield
When someone bakes a cake and starts with 100g of ingredients, the final cake is likely to have a mass of less than 100g (Figure 3.12). This will be because some of the ingredients will be lost on the bowl, on the spoon or on the eclair, and some crumbs of cake will be lost when it is taken out of the cake tin.

In a similar way, when carrying out chemical reactions we are unlikely to produce all that we expect to. There are many reasons for this:

1. Some reactions do not go to completion (i.e. they do not completely finish) – sometimes this is because they are irreversible and some of the products may turn back into reactants.

2. Some of the product may be lost when it is separated from the reaction mixture – for example, some may be left on the apparatus.

3. Some of the reactants may react in ways different to the desired reaction – in other words some of the reactants may take part in other reactions as well.

The amount of product made in a reaction is called the yield. The percentage yield is a measure of the amount produced in a reaction compared to the maximum theoretical amount that is expected as a percentage. For example, if a reaction was expected to form a maximum theoretical 20g of product but only 10g was made, then the yield is 10g and the percentage yield is 50%.

Example
In a reaction the maximum theoretical mass of product was 45g, the yield produced was 19g. Calculate the percentage yield.

Answer
Percentage yield = \( \frac{\text{mass of product actually made}}{\text{maximum theoretical mass of products}} \times 100 \% = \frac{19g}{45g} \times 100 \% = 42.22 \% \)
**Atom economy**

Chemical reactions are used to make specific products that we want. In some reactions, there is only one product and all the atoms in the reactants end up in that product, for example, when ethanol reacts with steam to make ethylene (C2H4O)H, all of the atoms in the reactants end up in the desired product ethanol.

\[
\text{C}_2\text{H}_6\text{O} + \text{H}_2\text{O} \rightarrow \text{C}_2\text{H}_4\text{O} + \text{H}_2\text{O}
\]

However, in many reactions, there are one or more other products as well as the desired product. In these reactions, not all of the atoms that start with end up in the desired product. For example, when ethanol is made by fermentation from glucose, carbon dioxide is formed as well as the ethanol. This means that some of the atoms from the glucose do not end up in the ethanol.

\[
\text{C}_6\text{H}_{12}\text{O}_6 \rightarrow 2\text{C}_2\text{H}_4\text{O} + 2\text{CO}_2
\]

The atom economy is a way of measuring what percentage of the mass of all the atoms in the reactants ends up in the desired product.

**Example**

Calculate the atom economy when ethanol is made from glucose by this reaction. Give your answer to 3 significant figures.

\[
\text{C}_2\text{H}_6\text{O} + \text{H}_2\text{O} \rightarrow \text{C}_2\text{H}_4\text{O} + \text{H}_2\text{O}
\]

**Answer**

Ethanol 

\[
\begin{align*}
\text{C}_2\text{H}_4\text{O} + \text{H}_2\text{O} & \rightarrow \text{C}_2\text{H}_4\text{O} + \text{H}_2\text{O} \\
\text{C}_2\text{H}_4\text{O} + \text{H}_2\text{O} & \rightarrow \text{C}_2\text{H}_4\text{O} + \text{H}_2\text{O} \\
\text{Total mass of reactants} & = 30 \text{ g} \\
\text{Total mass of products} & = 30 \text{ g}
\end{align*}
\]

The atom economy for this reaction is 100%.

**Example**

Calculate the atom economy when titanium is made from titanium oxide by this reaction. Give your answer to 2 significant figures.

\[
\text{Ti}_2\text{O}_3 + 3\text{H}_2 \rightarrow 2\text{Ti} + 3\text{H}_2\text{O}
\]

**Answer**

Titanium oxide 

\[
\begin{align*}
\text{Ti}_2\text{O}_3 + 3\text{H}_2 & \rightarrow 2\text{Ti} + 3\text{H}_2\text{O} \\
\text{Total mass of reactants} & = 156 \text{ g} \\
\text{Total mass of products} & = 156 \text{ g}
\end{align*}
\]

The atom economy for this reaction is 100%.

**Example**

Calculate the atom economy when water is made from hydrogen and oxygen by this reaction. Give your answer to 2 significant figures.

\[
\text{2H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}
\]

**Answer**

Hydrogen 

\[
\begin{align*}
\text{2H}_2 + \text{O}_2 & \rightarrow 2\text{H}_2\text{O} \\
\text{Total mass of reactants} & = 2 \times 2 = 4 \text{ g} \\
\text{Total mass of products} & = 2 \times 2 = 4 \text{ g}
\end{align*}
\]

The atom economy for this reaction is 100%.

The higher the atom economy of a reaction, the more of the mass of the reactants that ends up in the desired product. This means that there is less waste. For example, if 10 tonnes of reactants are used in a reaction with an atom economy of 80%, then 8 tonnes of the desired product are formed plus 2 tonnes of waste materials. However, if the same product could be made from 80% of starting materials in a reaction with an atom economy of 90%, then 4 tonnes of the desired product is formed along with only 2 tonnes of waste material.

As the Earth’s resources are precious, the less waste material produced in any process the better. In the long term, processes with higher atom economy are more sustainable. They can also be more economic as there is less waste to dispose of.
Test yourself

37. Calculate the mass of waste material in a reactor with 10% atom accuracy if 1 kg of material is used.
38. Calculate the mass of waste material in a reactor with 20% atom accuracy if 1 kg of material is used.
39. Calculate the mass of waste material in a reactor with 30% atom accuracy if 1 kg of material is used.
40. Explain why reactions with higher atom accuracies are more reliable.
41. Calculate the amount of material in a reactor with 40% atom accuracy if 1 kg of material is used.
42. Calculate the amount of material in a reactor with 50% atom accuracy if 1 kg of material is used.
43. Calculate the amount of material in a reactor with 60% atom accuracy if 1 kg of material is used.

Show you can...

60. Calculate the number of moles of different gases at the same temperature and pressure.

Volume (L)

Argon
Oxygen
Carbon dioxide

The volume of gases

The volume of a gas varies with temperature and pressure:
- The higher the temperature of a gas, the greater its volume.
- The greater the pressure of a gas, the smaller its volume.

However, providing the temperature and pressure of gases are the same, an equal number of moles of all gases have the same volume (Figure 3.13).

In other words, the volume of a gas does not depend on which gas it is.

Example

What is the volume of the following gases at room temperature and pressure?

Answer:
- 10 moles of O₂, volume = 24 L
- 5 moles of CO₂, volume = 12 L

Example

How many moles of each of the following gases whose volumes are measured at room temperature and pressure?

Answer:
- 48 L of CO₂, moles = 2 moles
- 12 L of O₂, moles = 0.5 moles

Figure 3.14: The same number of particles of different gases have the same volume at the same temperature and pressure.
Example

What is the volume of 1 g of methane gas at room temperature and pressure?

Answer

\[ \text{mass of CH}_4 \times \text{molar mass} = \text{volume at room temp and pressure} \]

Volume = 26 L

Example

What is the mass of 0.5 dm³ of hydrogen gas measured at room temperature and pressure?

Answer

\[ \text{volume of H}_2 \times \text{molar mass} = \text{mass at room temp and pressure} \]

Mass = 28 g

Test yourself

43. Calculate the volume of the following gases at room temperature and pressure:
   a. 3 moles of hydrogen (H₂)
   b. 4 moles of oxygen (O₂)
   c. 2 moles of nitrogen (N₂)
   d. 1 mole of carbon dioxide (CO₂)
   e. 5 moles of water (H₂O)

44. Calculate the number of moles in each of the following gases at room temperature and pressure. Give your answer to 3 significant figures.
   a. 25 g of carbon monoxide (CO)
   b. 20 g of ammonia (NH₃)
   c. 30 g of hydrogen (H₂)

45. Calculate the mass of each of the following gases at room temperature and pressure. Give your answer to 3 significant figures.
   a. 22 cm³ of oxygen (O₂)
   b. 15 cm³ of nitrogen (N₂)
   c. 40 cm³ of argon (Ar)

Show you can...

copy and complete the following table:

<table>
<thead>
<tr>
<th>Volume of gas</th>
<th>molar mass</th>
<th>volume at room temperature and pressure</th>
<th>number of moles</th>
<th>number of atoms</th>
</tr>
</thead>
<tbody>
<tr>
<td>5 dm³</td>
<td>28 g</td>
<td>26 L</td>
<td>2 moles</td>
<td>2 x 6.02 x 10²³</td>
</tr>
</tbody>
</table>

Example

What volume of oxygen reacts with 10 dm³ of hydrogen with the volume of both gases measured at the same temperature and pressure?

Answer

\[ \text{volume of oxygen} = \frac{10 \text{ dm}³ \times 2 \text{ moles of oxygen}}{5 \text{ moles of hydrogen}} \]

Volume of oxygen = 4 dm³

Example

What volume of carbon dioxide reacts with 40 dm³ of oxygen with the volume of both gases measured at the same temperature and pressure?

Answer

\[ \text{volume of carbon dioxide} = \frac{40 \text{ dm}³ \times 1 \text{ mole of carbon dioxide}}{5 \text{ moles of oxygen}} \]

Volume of carbon dioxide = 8 dm³

Example

Due to equal amounts of moles of different gases having the same volume (at the same temperature and pressure), we can work out the volume of gases involved in chemical reactions. Figures 3.15 and 3.16 show how the molar ratio from the equation can be used to do this.

Example

\[ \text{Carbon dioxide} + \text{hydrogen} \rightarrow \text{water} \]

\[ \text{2 moles of} \ \text{H}_2 \ \text{reacts with 1 mole of} \ \text{CO}_2 \]

Therefore, volume of CO₂ = 5 x volume of H₂ = 5 x 40 = 200 dm³
The concentration of solutions

**Concentration of solutions in g/dm$^3$**

Figure 3.17 shows two solutions of copper sulphate. The one that is darker blue has much more copper sulphate 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 stated in g/dm$^3$, which means the number of grams of solute dissolved in each dm$^3$ of solution. Figure 3.18, if B is the same volume as 1000 cm$^3$ or 1 dm$^3$, If 50 g of copper sulphate is dissolved in 1 dm$^3$ of solution, then the concentration is 50 g/dm$^3$.

In the laboratory, we often use volumes measured in cm$^3$ rather than dm$^3$. If there are 1000 cm$^3$ in 1 dm$^3$, we should divide the volume in cm$^3$ by 1000 to find the volume in dm$^3$. For example, 25 cm$^3$ is $25 \div 1000 = 0.025$ dm$^3$.

**Test yourself**

46. Calculate the volume of nitrogen that reacts with 10 dm$^3$ of hydrogen with the volume of both gases measured at the same temperature and pressure, $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

47. Calculate the volume of nitrogen oxide (NO) that is formed when 10 dm$^3$ of nitrogen monoxide ($NO\_2$) dissolves with the volume of nitric oxide ($NO$) measured at the same temperature and pressure, $NO\_2 + NO = 2NO_2$.

48. Calculate the volume of oxygen that reacts with 10 dm$^3$ of carbon with the volume of both gases measured at the same temperature and pressure, $C(s) + O_2(g) \rightarrow CO_2(g)$.

49. Calculate the volume of argon that reacts with 10 dm$^3$ of hydrogen with the volume of both gases measured at the same temperature and pressure, $2H_2(g) + Ar(g) = 2H_2(g) + Ar(g)$.

**Examples**

Find the concentration of the following solutions in g/dm$^3$.

- 80 g dissolved in 4 dm$^3$: Concentration = volume / mass
- 75 g dissolved in 20 dm$^3$: $C = \frac{75}{20} = 3.75$ g/dm$^3$
- 150 g dissolved in 10 dm$^3$: $C = \frac{150}{10} = 15$ g/dm$^3$

**Concentration of solutions in mol/dm$^3$**

We can measure the concentration of a solution in moles/dm$^3$. This is effectively the number of moles of solute dissolved in each dm$^3$ of solution. Figure 3.19B, if A is the same volume as 1000 cm$^3$ or 1 dm$^3$. For example, if 2 moles of HCl are dissolved in 1 dm$^3$ of solution, then the concentration is 2 mol/dm$^3$.

**Test yourself**

50. What is the concentration of copper sulphate in a solution in which 100 g of solute are dissolved in 2 dm$^3$ of solution?

51. What is the concentration of copper in a solution in which 10 g of solute are dissolved in 0.5 dm$^3$ of solution?

52. What is the concentration in moles/dm$^3$ of a solution in which 2 kg of solute are dissolved in 10 dm$^3$ of solution?

53. What mass of solute is dissolved in 10 dm$^3$ of a solution with concentration 2 mol/dm$^3$?

54. What mass of solute is dissolved in 35 cm$^3$ of a solution with concentration 3 mol/dm$^3$?

**Example**

Find the concentration of each of the following solutions in moles/dm$^3$.

- 0.01 mol dissolved in 1 dm$^3$: concentration = moles / dm$^3$ = $0.01 / 1 = 0.01$ mol/dm$^3$
- 0.02 mol dissolved in 0.5 dm$^3$: $C = \frac{0.02}{0.5} = 0.04$ mol/dm$^3$
- 0.05 mol dissolved in 2 dm$^3$: $C = \frac{0.05}{2} = 0.025$ mol/dm$^3$
**Example**

And the number of moles of solute in each of the following solutions.

<table>
<thead>
<tr>
<th>Solution</th>
<th>Moles of solute</th>
<th>Concentration</th>
</tr>
</thead>
<tbody>
<tr>
<td>2.5 M NaCl</td>
<td>2.5 moles</td>
<td>2.5 M</td>
</tr>
<tr>
<td>0.01 M HCl</td>
<td>0.01 moles</td>
<td>0.01 M</td>
</tr>
</tbody>
</table>

**Test yourself**

1. Calculate the concentration of the following solutions in molarity:
   - a) 0.1 moles dissolved in 0.1 L of solution
   - b) 0.25 moles dissolved in 0.5 L of solution
   - c) 0.05 moles dissolved in 0.1 L of solution
   - d) 0.01 moles dissolved in 0.05 L of solution

2. Calculate the number of moles of solute in the following solutions:
   - a) 0.1 M NaCl dissolved in 0.5 L of solution
   - b) 0.05 M CaCl₂ dissolved in 0.25 L of solution

3. Calculate the concentration of the following solutions in molarity:
   - a) 0.1 M H₂SO₄ dissolved in 0.2 L of solution
   - b) 0.01 M FeCl₃ dissolved in 0.01 L of solution

4. We can easily convert concentrations in mol/L to g/dm³ using this equation:

   \[ \text{concentration (g/dm³)} = \frac{M \times \text{concentration (mol/L)}}{1000} \]

**Example**

What is the concentration in grams of a solution of concentration 0.5 M NaCl?

**Answer:**

At 0.5 M NaCl = 0.5

Concentration (g/dm³) = \[ \frac{1000 \times 0.5}{0.5} \] = 1000 g/dm³

**Show you can...**

A solution of some acid and base is neutralised in solution. What is the concentration of the acid or base and why?

**Titrations**

Titrations are a very accurate experimental technique that can be used to find the concentration of a solution by reacting it with a solution of known concentration. Titrations are often used to find the concentration of acids or bases.

Titrations are apparatus including a pipette, conical flask and burette (Figure 3.20). A pipette is a glass tube designed to measure a specific volume of a solution very accurately. A typical pipette measures out 25 cm³ within a margin of ±0.01 cm³. The pipette is filled using a pipette filler which is attached to the end of the pipette. A burette is a glass tube with a tap (to let out the liquid) with markings on to show the volume to the nearest 0.1 cm³.

The following steps are followed in a titration:

1. A known volume of a solution of an acid or alkali is measured out using a pipette and placed into a conical flask.
2. A few drops of a suitable indicator are added. For most acid-base titrations, methyl orange or phenolphthalein is suitable.
3. The other solution, the acid or alkali, is added to the conical flask from a burette.
4. The solution is added from the burette until the indicator changes colour (the end point). The solution is added dropwise around the point where the indicator changes colour to ensure the exact volume required is used.
5. The volume added from the burette is recorded.
6. The experiment is repeated until consistent results are achieved (i.e., results that are very close to each other). The mean volume is found using the consistent results.

Titrations are a very accurate technique and when done correctly often give a value within 3% of the true value.

**Titrations calculation**

The concentration of a solution can be found using the results of a titration. The volume of both solutions has been measured and the concentration of one of the solutions will have been known at the start. This information is used as shown below.

**1** Work out the moles of the solution whose concentration is known (using moles = concentration (concentrate / volume (dm³)))

**2** Use the moles of the other solution, with that of a known solution, in the balanced equation to work out the moles of that solution.

**3** Calculate the concentration of the other solution (using concentration = moles / volume (dm³))
### Example

#### 1. Determination of the concentration of a solution using titration

A solution of NaOH was standardized using the titration method. The titration was performed with a standard solution of HCl. The reaction is:

\[ \text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O} \]

**Questions**

1. Calculate the concentration of the NaOH solution used in the titration.
2. Calculate the volume of HCl solution used.
3. Determine the molarity of the HCl solution.

#### 2. Determination of the concentration of a solution using titration

A solution of NH₃ was standardized using the titration method. The titration was performed with a standard solution of HCl. The reaction is:

\[ \text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4\text{Cl} \]

**Questions**

1. Calculate the concentration of the NH₃ solution used in the titration.
2. Calculate the volume of HCl solution used.
3. Determine the molarity of the HCl solution.

### Test yourself

#### 1. Acid-Base titrations

A solution of HCl was titrated with a solution of NaOH. The reaction is:

\[ \text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O} \]

**Questions**

1. Calculate the concentration of the HCl solution used in the titration.
2. Calculate the volume of NaOH solution used.
3. Determine the molarity of the NaOH solution.

#### 2. Oxidation-Reduction titrations

A solution of Fe⁺₂⁺ was titrated with a solution of KMnO₄. The reaction is:

\[ \text{Fe}⁺₂⁺ + \text{KMnO}_4 + \text{H}_2\text{SO}_4 \rightarrow \text{Fe}³⁺ + \text{Mn}²⁺ + \text{K}_2\text{SO}_4 + \text{H}_2\text{O} \]

**Questions**

1. Calculate the concentration of the KMnO₄ solution used in the titration.
2. Calculate the volume of Fe⁺₂⁺ solution used.
3. Determine the molarity of the Fe⁺₂⁺ solution.

### Show you can...:

1. Describe the titration process and its importance.
2. Explain the calculation of the concentration from a titration.
3. Perform a titration experiment.

### Determination of the reacting volumes of solutions of a strong acid and a strong base by titration

A solution of a strong acid (H₂SO₄) was titrated with a solution of a strong base (NaOH). The reaction is:

\[ \text{H}_2\text{SO}_4 + \text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O} \]

**Questions**

1. Calculate the volume of NaOH solution used in the titration.
2. Calculate the concentration of the H₂SO₄ solution.
3. Determine the molarity of the NaOH solution.

### Further reading

Another section of this protocol is found on page 112. In Chapter 4.
Chapter review questions

1. Calculate the relative formula mass (M_r) of the following substances:
   a. SnO₂
   b. Na₂S
   c. Ca(NO₃)₂
   d. NH₄Cl
   e. MgCl₂
   f. Fe(OH)₃

2. A 100 g sample of iron reacts with 50 g of oxygen to form iron oxide. What mass of iron oxide is formed?

3. What happens in a reaction where calcium carbonate is heated to decompose in a reaction 2CaCO₃ (s) → CaO (s) + CO₂ (g)?

4. In a reaction, 11.2 g of copper reacts with 8.0 g of sulfur to form copper sulfide. What mass of copper sulfide is formed?

5. When sodium carbonate is heated, it decomposes in a reaction Na₂CO₃ (s) → Na₂O (s) + CO₂ (g). What mass of sodium oxide is formed?

6. When calcium carbonate is heated, it decomposes in a reaction 2CaCO₃ (s) → CaO (s) + CO₂ (g). What mass of calcium oxide is formed?

7. What is the mass of a mole of the following substances?
   a. potassium oxide, K₂O
   b. magnesium oxide, MgO
   c. iron (III) oxide, Fe₂O₃
   d. iron (II) oxide, FeO

8. What is the mass of each of the following?
   a. 2 moles of calcium oxide, CaO
   b. 3 moles of potassium oxide, K₂O
   c. 4 moles of copper (II) oxide, CuO

9. How many moles of potassium oxide would be needed to react with 5 moles of iron (III) oxide?

10. How many moles of calcium oxide would be needed to react with 2 moles of potassium hydroxide?

11. How many moles of sodium chloride would be needed to react with 3 moles of sodium hydroxide?

12. What volume of hydrogen gas is formed, measured at room temperature and pressure when 2.5 g of sodium reacts with water?

13. What volume of carbon dioxide gas, measured at room temperature and pressure, is released when 100 g of sodium carbonate reacts with 100 g of 5% sulfuric acid solution?

14. What volume of carbon dioxide gas, measured at room temperature and pressure, is released when 100 g of sodium carbonate reacts with 100 g of 2% hydrochloric acid solution? (One of the reactants is in excess.)
Practice questions

1. Which one of the following would contain the same number of moles as 4.2 g of magnesium? (1 mark)
   a) 34 g of carbon
   b) 84 g of carbon
   c) 64 g of carbon
   d) 44 g of carbon

2. What is the mass of one mole of calcium oxide (CaO)? (1 mark)
   a) 44 g
   b) 44 kg
   c) 44 mg
   d) 44 g/cm³

3. Most metals are found naturally in rocks called ores. Some examples are shown in the table.

<table>
<thead>
<tr>
<th>Mineral</th>
<th>Formula or Compound Used</th>
</tr>
</thead>
<tbody>
<tr>
<td>Chalcopyrite</td>
<td>CuFeS₂</td>
</tr>
<tr>
<td>Galena</td>
<td>PbS</td>
</tr>
<tr>
<td>Pyrite</td>
<td>Fe₂S₃</td>
</tr>
</tbody>
</table>

What is the mass of one mole of copper oxide (CuO)? (1 mark)

4. A cup of water is poured into a beaker containing 300 mL of a solution. The mass of water is then measured with an accuracy of 0.1 g. If the solution has a density of 1.2 g/mL, what is the mass of water poured into the beaker? (2 marks)

5. The volume of a gas is 450 mL at 25°C and 1 atm pressure. If the temperature is increased to 35°C at constant pressure, what is the new volume of the gas? (2 marks)

6. A solution contains 0.5 moles of NaCl. If the solution is diluted to 2 liters, what is the new concentration of NaCl? (2 marks)

7. A student measures the mass of a small sample of copper to be 5.5 g. If the density of copper is 8.96 g/cm³, what is the volume of the copper sample? (2 marks)

8. A 250 mL beaker contains 200 mL of water and 50 mL of a solution. If the solution is 10% by mass sodium chloride, what is the mass of sodium chloride in the solution? (2 marks)

9. A student conducts an experiment to determine the density of a block of aluminum. The block has a mass of 50 g and a volume of 10 cm³. What is the density of the aluminum? (2 marks)

10. A student conducts an experiment to determine the percentage of sodium chloride in a solution. The mass of the sample is 20 g and the mass of the precipitate formed is 2 g. What is the percentage of sodium chloride in the solution? (2 marks)

11. A student conducts an experiment to determine the percentage of iron in a sample of rust. The mass of the sample is 100 g and the mass of the rust is 50 g. What is the percentage of iron in the rust? (2 marks)

12. A student conducts an experiment to determine the percentage of copper in a sample of copper sulfate. The mass of the sample is 50 g and the mass of copper is 20 g. What is the percentage of copper in the copper sulfate? (2 marks)

13. A student conducts an experiment to determine the percentage of calcium in a sample of calcium carbonate. The mass of the sample is 100 g and the mass of calcium carbonate is 50 g. What is the percentage of calcium in the calcium carbonate? (2 marks)

14. A student conducts an experiment to determine the percentage of magnesium in a sample of magnesium sulfate. The mass of the sample is 100 g and the mass of magnesium sulfate is 50 g. What is the percentage of magnesium in the magnesium sulfate? (2 marks)

15. A student conducts an experiment to determine the percentage of potassium in a sample of potassium chloride. The mass of the sample is 100 g and the mass of potassium chloride is 50 g. What is the percentage of potassium in the potassium chloride? (2 marks)

16. A student conducts an experiment to determine the percentage of sodium in a sample of sodium chloride. The mass of the sample is 100 g and the mass of sodium chloride is 50 g. What is the percentage of sodium in the sodium chloride? (2 marks)

17. A student conducts an experiment to determine the percentage of hydrogen in a sample of water. The mass of the sample is 100 g and the mass of hydrogen is 10 g. What is the percentage of hydrogen in the water? (2 marks)

18. A student conducts an experiment to determine the percentage of oxygen in a sample of water. The mass of the sample is 100 g and the mass of oxygen is 80 g. What is the percentage of oxygen in the water? (2 marks)

19. A student conducts an experiment to determine the percentage of carbon in a sample of carbon dioxide. The mass of the sample is 100 g and the mass of carbon dioxide is 72 g. What is the percentage of carbon in the carbon dioxide? (2 marks)

20. A student conducts an experiment to determine the percentage of nitrogen in a sample of nitrogen dioxide. The mass of the sample is 100 g and the mass of nitrogen dioxide is 72 g. What is the percentage of nitrogen in the nitrogen dioxide? (2 marks)

21. A student conducts an experiment to determine the percentage of argon in a sample of argon gas. The mass of the sample is 100 g and the mass of argon gas is 40 g. What is the percentage of argon in the argon gas? (2 marks)

22. A student conducts an experiment to determine the percentage of helium in a sample of helium gas. The mass of the sample is 100 g and the mass of helium gas is 4 g. What is the percentage of helium in the helium gas? (2 marks)

23. A student conducts an experiment to determine the percentage of oxygen in a sample of water. The mass of the sample is 100 g and the mass of oxygen is 16 g. What is the percentage of oxygen in the water? (2 marks)

24. A student conducts an experiment to determine the percentage of hydrogen in a sample of water. The mass of the sample is 100 g and the mass of hydrogen is 8 g. What is the percentage of hydrogen in the water? (2 marks)

25. A student conducts an experiment to determine the percentage of nitrogen in a sample of nitrogen gas. The mass of the sample is 100 g and the mass of nitrogen gas is 28 g. What is the percentage of nitrogen in the nitrogen gas? (2 marks)

26. A student conducts an experiment to determine the percentage of argon in a sample of argon gas. The mass of the sample is 100 g and the mass of argon gas is 40 g. What is the percentage of argon in the argon gas? (2 marks)

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47. A student conducts an experiment to determine the percentage of helium in a sample of helium gas. The mass of the sample is 100 g and the mass of helium gas is 4 g. What is the percentage of helium in the helium gas? (2 marks)
Working scientifically: Interconverting units

Notes

Keep all the calculations used in chemistry will require different units. It is important that you can convert between units.


Example

What is 15 cm in dm?

Answer

To convert from cm to dm, you need to divide by 100.

\(1.5 \text{ dm} = 15 \div 100 = 0.15 \text{ dm}\)

Example

What is 0.4 km in m?

Answer

To convert from km to m, you need to multiply by 1000.

\(400 \text{ m} = 0.4 \times 1000 = 400 \text{ m}\)

Notes

Mass can be measured in: milligrams (mg), grams (g), kilograms (kg) and tonnes.

- 1 tonne = 1000 kg
- 1 kilogram = 1000 g
- 1 gram = 1000 mg

The flow diagram in Figure 2.21 will help you to convert between mass units.

\[\begin{align*}
\text{tonne} & \quad \times 1000 \\
\text{kilogram} & \quad \times 1000 \\
\text{gram} & \quad \times 1000 \\
\text{milligram} & \quad \times 1000
\end{align*}\]

Example

Convert 600 mg to grams.

Answer

To convert from mg to g, you need to divide by 1000.

\(600 \text{ mg} = 0.6 \text{ g}\)

Example

Convert 3.2 kg to grams.

Answer

To convert from kg to g, you need to multiply by 1000.

\(3200 \text{ g} = 3.2 \times 1000 = 3200 \text{ g}\)

Example

Convert 0.4 km to m.

Answer

To convert from km to m, you need to multiply by 1000.

\(400 \text{ m} = 0.4 \times 1000 = 400 \text{ m}\)

Example

Convert 250 mg to kg.

Answer

First, convert mg to g by dividing by 1000.

\(0.25 \text{ g} = 250 \div 1000 = 0.25 \text{ g}\)

Then convert g to kg by dividing by 1000.

\(0.00025 \text{ kg} = 0.25 \div 1000 = 0.00025 \text{ kg}\)

Questions

1. Convert the following volumes in the units shown:
   a. 25 cm³ to dm³
   b. 1500 cm³ to dm³
   c. 10 ml to dm³
   d. 2 l to dm³

2. Carry out the following unit conversions:
   a. 1 mm to m
   b. 1000 mg to g
   c. 3 kg to g
   d. 2 l to ml

Tips

Think carefully when converting between units. A kilogram is bigger than a gram so when converting from kilogram to grams you would expect to get a bigger number.
4 Chemical changes

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.

Previously you could have learned:
- Metals can be listed in a reactivity series which compares metals in terms of their reactivity.
- Metals are extracted from compounds in ore.
- Metals with high 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
- Name two reactive metals.
- Name two non-reactive metals.
- Name the metal oxide which contains sodium (Na) and chlorine (Cl).
- How is sodium extracted from this metal oxide?
- What gas is made when acids react with:
  - metals
  - carbonates?
- State whether each of the following solutions is acidic, neutral or alkaline:
  - pH 1
  - pH 7
  - pH 2
  - pH 9

The reactivity series of metals

Metals have many uses. For example, they are used in electrical cables, cars, airplanes, 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 4.3 shows some common metals. Carbon and hydrogen are included for comparison although they are non-metals.
Reaction with oxygen
Most metals react with oxygen. Reactive metals burn when they are heated, as they react with oxygen. In the table, the metals in the middle of the reactivity series react with oxygen when heated and can burn if the metal is powdered. Copper, a less reactive metal, reacts with oxygen forming a layer of copper oxide on the surface of the copper but does not burn (Figure 4.2). 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:
metal + oxygen → metal oxide

For example:
sodium + oxygen → sodium oxide
4Na + O₂ → 2Na₂O
copper + oxygen → copper oxide
2Cu + O₂ → 2CuO

These are examples of oxidisation reactions. An oxidisation 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 4.3).

Metal + water → metal hydroxide + hydrogen

Table 4.1 shows how metals react with cold water.

<table>
<thead>
<tr>
<th>Metal</th>
<th>Reaction with water</th>
</tr>
</thead>
<tbody>
<tr>
<td>Potassium</td>
<td>reacts violently and gives off hydrogen gas</td>
</tr>
<tr>
<td>Sodium</td>
<td>reacts violently and gives off hydrogen gas</td>
</tr>
<tr>
<td>Lithium</td>
<td>reacts violently and gives off hydrogen gas</td>
</tr>
<tr>
<td>Calcium</td>
<td>reacts very slowly</td>
</tr>
<tr>
<td>Magnesium</td>
<td>reacts slowly</td>
</tr>
<tr>
<td>Aluminum</td>
<td>reacts very slowly</td>
</tr>
<tr>
<td>Zinc</td>
<td>reacts slowly</td>
</tr>
<tr>
<td>Iron</td>
<td>reacts slowly</td>
</tr>
<tr>
<td>Copper</td>
<td>no reaction</td>
</tr>
<tr>
<td>Silver</td>
<td>no reaction</td>
</tr>
<tr>
<td>Gold</td>
<td>no reaction</td>
</tr>
<tr>
<td>Platinum</td>
<td>no reaction</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.
metal + acid → metal salt + 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 quickly. Metals that are less reactive than hydrogen do not react with dilute acids.

Table 4.2 shows how metals react with dilute hydrochloric acid.

<table>
<thead>
<tr>
<th>Metal</th>
<th>Reaction with dilute hydrochloric acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>Potassium</td>
<td>reacts violently</td>
</tr>
<tr>
<td>Sodium</td>
<td>reacts violently</td>
</tr>
<tr>
<td>Lithium</td>
<td>reacts violently</td>
</tr>
<tr>
<td>Calcium</td>
<td>reacts very slowly</td>
</tr>
<tr>
<td>Magnesium</td>
<td>reacts very slowly</td>
</tr>
<tr>
<td>Aluminum</td>
<td>reacts very slowly</td>
</tr>
<tr>
<td>Zinc</td>
<td>reacts very slowly</td>
</tr>
<tr>
<td>Copper</td>
<td>no reaction</td>
</tr>
<tr>
<td>Silver</td>
<td>no reaction</td>
</tr>
<tr>
<td>Gold</td>
<td>no reaction</td>
</tr>
<tr>
<td>Platinum</td>
<td>no reaction</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(III) oxide because aluminium is more reactive than iron.

\[ \text{aluminium} + \text{iron(III) oxide} \rightarrow \text{aluminium oxide} + \text{iron} \]

This reaction is used to weld railways (Figure 4.5). A mixture of aluminium and iron oxide is placed over the gap between the railway lines and the reaction starts. The reaction gets very hot and produces molten iron which flows into the gap, seals 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 basic and so the solution turns blue as the copper nitrate is formed (Figure 4.4).

\[ \text{copper} + \text{silver nitrate} \rightarrow \text{copper nitrate} + \text{silver} \]

\[ Cu + 2AuNO₃ → Cu(NO₃)₂ + 2Ag \]

**Test yourself**

1. Complete the following word equations, or write no reaction if the reaction would not take place:

   - calcium + oxygen
   - copper + water
   - calcium + nitric acid
   - copper + hydrochloric acid
   - Fe + magnesium nitrate
   - Zn + magnesium sulphate
   - Mg + H₂O
   - Cu + H₂SO₄
   - Zn + H₂SO₄
   - Zn + CuSO₄
   - Mg + CuSO₄
   - Mg + H₂SO₄
   - Zn + H₂SO₄
   - Zn + CuSO₄
   - Mg + CuSO₄
   - Cu + H₂SO₄
   - Zn + H₂SO₄
   - Zn + CuSO₄

2. Write a word equation for the following reactions:

   - reaction of calcium with water

3. What happens when calcium reacts with water?

4. Calculate the reaction of calcium with water.

5. Write a balanced chemical equation for the reaction of magnesium with water.

6. What happens when magnesium reacts with water?

7. Write a word equation for the reaction of copper with water.

8. What happens when copper reacts with water?

9. Write a balanced chemical equation for the reaction of zinc with acid.

10. What happens when zinc reacts with acid?

**Show you can...**

- Determine the order of reactivity of the metals copper, magnesium, and zinc by reacting each metal with the oxides of other metals and observing the results obtained (answer to the related questions).

- From the list below, write word and balanced chemical equations for the following reactions:

   - copper oxide + hydrochloric acid
   - copper oxide + nitric acid
   - magnesium oxide + hydrochloric acid
   - magnesium oxide + nitric acid
   - copper oxide + hydrochloric acid
   - copper oxide + nitric acid
   - magnesium oxide + hydrochloric acid
   - magnesium oxide + nitric acid
   - copper oxide + hydrochloric acid
   - copper oxide + nitric acid
   - magnesium oxide + hydrochloric acid
   - magnesium oxide + nitric acid

- Write a balanced chemical equation for the reaction of hydrochloric acid with copper oxide.

- Write a balanced chemical equation for the reaction of nitric acid with magnesium oxide.

- Write a balanced chemical equation for the reaction of hydrochloric acid with zinc oxide.

- Write a balanced chemical equation for the reaction of nitric acid with zinc oxide.
Oxidation and reduction in terms of electrons

Oxidation can be defined as a reaction where a substance gains oxygen, whereas 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.

One way to remember this is the phrase OIL RIG (Figure 4.7).

When metals react with oxygen, water and acids, the metal atoms lose electrons and form metal ions (Table 4.3). This means that each reaction can be defined as oxidation in terms of the loss of electrons. However, the reaction with oxygen is the only one that can be defined as oxidation in terms of gaining oxygen.

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 aluminum (Figure 4.4).

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 (oxidation-reduction).

Writing ionic equations and/or half equations for displacement reactions

Ionic equations and/or half equations can be written for displacement reactions. (See pages 69-70 for more help on writing these equations.)

Example

Write two half equations for the displacement of zinc from zinc oxide by iron.

- Oxidation: \( \text{ZnO} \rightarrow \text{Zn}^{2+} + 2\text{e}^- \)
- Reduction: \( \text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn} \)

Answer

In this reaction, the Zn ions become Zn atoms, while the Fe ions become Fe atoms.

The ionic equation for this reaction is:

\[ \text{ZnO} + \text{Fe} \rightarrow \text{Zn} + \text{FeO} \]

Example

Write an ionic equation and two half equations for the displacement of silver from silver nitrate by copper.

- Oxidation: \( \text{Ag}^{+} + \text{e}^- \rightarrow \text{Ag} \)
- Reduction: \( \text{Cu} + \text{Ag}^{+} \rightarrow \text{Cu}^{2+} + \text{Ag} \)

Answer

In this reaction, the Cu ions become Cu atoms, while the Ag ions become Ag atoms.

The ionic equation for this reaction is:

\[ \text{AgNO}_3 + \text{Cu} \rightarrow \text{Cu(NO}_3)_2 + \text{Ag} \]

Example

Write an ionic equation and two half equations for the displacement of copper from copper sulfate by iron.

- Oxidation: \( \text{Fe} + \text{Cu}^{2+} \rightarrow \text{Fe}^{2+} + \text{Cu} \)
- Reduction: \( \text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu} \)

Answer

In this reaction, the Fe atoms become Fe ions, while the Cu ions become Cu atoms. We can leave out the CuO ions from the ionic equation as they do not change.

The ionic equation for this reaction is:

\[ \text{Fe} + \text{CuSO}_4 \rightarrow \text{FeSO}_4 + \text{Cu} \]
**Extraction of metals**

1. **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.

2. **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 4.11).

- **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 to iron this reaction (Figure 4.12).

- **Metals that are more reactive than carbon can be extracted by electrolysis.** This is studied in detail on page 137.

When metals are extracted from compounds, the metal ions in the compound gain electrons. For example, when iron is extracted from iron oxide, $Fe^{3+}$ ions in the iron oxide gain electrons to form $Fe^{2+}$ ions. This is reduction because the iron oxide loses oxygen and also because the $Fe^{2+}$ ions in the iron oxide gain electrons.

---

**Test yourself**

- **Section 1:** A metal in Group 1 of the periodic table is extracted vigorously with water.
  - Write a balanced equation for this reaction.
  - Write a balanced equation for this reaction in terms of ions.
  - Write the physical change in the reaction.
  - Write the chemical change in the reaction.
  - Write the overall equation for the reaction.

- **Section 2:** Magnesium reacts with copper oxide to produce magnesium oxide and copper.
  - Write a balanced equation for this reaction.
  - Write a balanced equation for this reaction in terms of ions.
  - Write the physical change in the reaction.
  - Write the chemical change in the reaction.
  - Write the overall equation for the reaction.

- **Section 3:** Magnesium reacts with copper oxide to produce magnesium oxide and copper.
  - Write a balanced equation for this reaction.
  - Write a balanced equation for this reaction in terms of ions.
  - Write the physical change in the reaction.
  - Write the chemical change in the reaction.
  - Write the overall equation for the reaction.

- **Section 4:** Examine the structure of Group 1 elements.
  - Explain the reactivity of Group 1 elements.
  - Explain why Group 1 elements react with oxygen.
  - Explain why Group 1 elements react with other metals.
  - Explain why Group 1 elements react with non-metals.

---

**Summary**

- **Extraction of metals**
  - There are many different methods of extracting metals.
  - The most common method is by heating the metal oxide with carbon or by electrolysis.
  - Many metals are found in nature, but most are found in compounds.
  - The metal can be extracted by heating the compound or by electrolysis.

---

**KEY TERMS**

- **Electrolysis**
  - The process of breaking down a compound into its constituent elements by passing an electric current through the compound.

---

**Figure 4.11**

- **Aluminium oxide:** The compound from which aluminium is extracted.
  - **Aluminium:** The metal obtained from aluminium oxide.

---

**Figure 4.12**

- **Iron oxide:** The compound from which iron is extracted.
  - **Iron:** The metal obtained from iron oxide.

---

**Figure 4.13**

- **Copper oxide:** The compound from which copper is extracted.
  - **Copper:** The metal obtained from copper oxide.

---

**Figure 4.14**

- **Aluminium oxide:** The compound from which aluminium is extracted.
  - **Aluminium:** The metal obtained from aluminium oxide.

---

**Figure 4.15**

- **Iron oxide:** The compound from which iron is extracted.
  - **Iron:** The metal obtained from iron oxide.

---

**Figure 4.16**

- **Copper oxide:** The compound from which copper is extracted.
  - **Copper:** The metal obtained from copper oxide.
Reactions of acids

What are acids and alkalis?

An acid is a substance that produces hydrogen ions, $H^+$, in aqueous solution. For example, solutions of:

- hydrochloric acid ($HCl$) contains hydrogen ($H^+$) ions and chloride ($Cl^-$) ions.
- sulfuric acid ($H_2SO_4$) contains hydrogen ($H^+$) ions and sulfate ($SO_4^{2-}$) ions.
- nitric acid ($HNO_3$) contains hydrogen ($H^+$) ions and nitrate ($NO_3^-$) ions.

An alkali is a substance that produces hydroxide ions, $OH^-$, in aqueous solution. For example, solutions of:

- sodium hydroxide ($NaOH$) contains sodium ($Na^+$) ions and hydroxide ($OH^-$) ions.
- potassium hydroxide ($KOH$) contains potassium ($K^+$) ions and hydroxide ($OH^-$) ions.
- calcium hydroxide ($Ca(OH)_2$) contains calcium ($Ca^{2+}$) ions and hydroxide ($OH^-$) 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 4.13).

Test yourself 11

11. What are acids and alkalis?
   a. A solution that produces hydrogen ions, $H^+$, in aqueous solution.
   b. A solution that produces hydroxide ions, $OH^-$, in aqueous solution.

Test yourself 12

12. What are acids and alkalis?
   a. A solution that produces hydrogen ions, $H^+$, in aqueous solution.
   b. A solution that produces hydroxide ions, $OH^-$, in aqueous solution.

Test yourself 13

13. What is the pH of a solution?
   a. A measure of the concentration of hydrogen ions in the solution.
   b. A measure of the concentration of hydroxide ions in the solution.

Test yourself 14

14. What is the pH of a solution?
   a. A measure of the concentration of hydrogen ions in the solution.
   b. A measure of the concentration of hydroxide ions in the solution.

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 4.14).
Strong and weak acids

Hydrochloric acid (HCl), hydrochloric sulfuric acid (H₂SO₄), and hydrochloric nitric acid (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 of their molecules break into ions in water (Figure 6.15).

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 6.17 shows the difference between the strong acid HCl and a weak acid H₂SO₄ in water.

If solutions 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. Citric acid in oranges, citric acid in citrus fruits (Figure 6.18), and carbonic acid in fizzy drinks are examples of weak acids in food and drink.

Example

Compare the following two solutions of acids:

1. A solution of a strong acid and a weak acid.
2. A solution of a strong acid and a strong acid.
3. A solution of a weak acid and a weak acid.

Aims

1. To compare the strength of acids.
2. To understand the difference between strong and weak acids.
3. To investigate the pH of solutions of different acids.

Test yourself

1. Which is a strong acid? Explain your answer.
2. Which is a weak acid? Explain your answer.
3. Which is the weakest acid? Explain your answer.

Answers

1. Explain and label the difference between these two solutions.
2. What is the difference between these two solutions?
3. What is the difference between these two solutions?
**Reaction of acids with metals**

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 forms as hydrogen gas is produced.

\[
\text{metal} + \text{acid} \rightarrow \text{metal salt} + \text{hydrogen}
\]

Table 4.4 shows how different acids from different types of salts. Table 4.5 shows how magnesium, zinc, and iron react with dilute hydrochloric and sulfuric acids.

When a metal reacts with an acid, a redox reaction takes place (Figure 4.39). The metal atoms lose electrons and so are oxidised. The ions in the acid gain electrons and are reduced.

**Reaction of acids with metal hydroxides**

Acids react with metal hydroxides to form a salt and water. Some examples of this reaction are shown in Table 4.6.

\[
\text{metal hydroxide} + \text{acid} \rightarrow \text{metal salt} + \text{water}
\]
Table 4.7

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbonic acid + water → carbonic acid + water</td>
<td>HCO₃⁻ + H₂O → H₂CO₃ + OH⁻</td>
</tr>
<tr>
<td>Copper (II) carbonate + water → copper (II) carbonate + water</td>
<td>CuCO₃(s) + H₂O(l) → CuCO₃(s) + H₂O(l)</td>
</tr>
</tbody>
</table>

In these reactions, the 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 forms as carbon dioxide gas is produced. Some examples of this reaction are shown in the Table 4.8.

Table 4.8

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Metal carbonate + acid → metal salt + water + carbon dioxide</td>
<td>MgCO₃(s) + 2HCl(aq) → MgCl₂(aq) + H₂O(l) + CO₂(g)</td>
</tr>
</tbody>
</table>

In these reactions, the ions from the acid are being used up and so they are examples of neutralisation reactions.

**Test yourself**

1. Complete the following word equations.
   a. Carbonic acid + hydrochloric acid
   b. Iron + sulfuric acid
   c. Barium hydroxide + hydrochloric acid
   d. Iron(II) hydroxide + hydrochloric acid
   e. Iron(II) carbonate + nitric acid
   f. Copper carbonate + hydrochloric acid

2. Write a balanced chemical equation for the reaction between iron and hydrochloric acid as shown in this equation: Fe(s) + 2HCl(aq) → FeCl₂(aq) + H₂(g). What does the symbol (aq) mean?

3. Write a balanced chemical equation for the reaction between copper(II) carbonate and hydrochloric acid. What does the symbol (s) mean?

4. Write a balanced chemical equation for the reaction between calcium hydroxide and sulfuric acid.

5. Write a balanced chemical equation for the reaction between magnesium hydroxide and hydrochloric acid.

**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 4.21)
- 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 ion reacts. The salt produced depends on which acid is used and what it reacts with. Table 4.9.1 below gives some examples of the salts formed when acids react.

Table 4.9

<table>
<thead>
<tr>
<th>Hydrochloric acid</th>
<th>Sodium carbonate</th>
<th>Sodium nitrate</th>
<th>Sodium chloride</th>
</tr>
</thead>
<tbody>
<tr>
<td>MgCl₂ + Na₂CO₃ → MgCO₃ + NaCl</td>
<td>MgCl₂ + Na₂CO₃ → MgCO₃ + 2NaCl</td>
<td>MgCl₂ + NaNO₃ → MgNO₃ + NaCl</td>
<td>MgCl₂ + NaCl → MgCl₂ + NaCl</td>
</tr>
</tbody>
</table>

**Show you can...**

- Sodium sulfooxydiphenyl is produced by the reaction between sulfuric acid and sodium hydroxide, to yield bismuthic nitrate.
- The percentage composition of sodium nitrate is shown in the table below.

<table>
<thead>
<tr>
<th>Sodium Nitrate</th>
<th>Sodium Sulfate</th>
<th>Copper Sulfate</th>
<th>Lead Sulfate</th>
</tr>
</thead>
<tbody>
<tr>
<td>78%</td>
<td>22%</td>
<td>5%</td>
<td>5%</td>
</tr>
</tbody>
</table>

Unique sodium was added to each of the products. Only one of the products reacted.
4 Chemical changes
Electrolysis

**KEY TERMS**
- Electrolyte: A liquid that conducts electricity.
- Electrolysis: The decomposition of ionic compounds using electricity.
- Anode: An electrode where oxidation takes place.
- Cathode: An electrode where reduction takes place.
- Electrolyte: A solution that conducts electricity.
- Electrolysis: The process of decomposing ionic compounds using electricity.

**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-metallic chloride.
- Copper sulfate (CuSO₄) – a combination of the metal copper with the non-metallic sulfate 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 positive ions are attracted to the negative electrode, and the negative ions are attracted to the positive electrode (Figure 4.62). 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.

**Electrolysis of molten ionic compounds**

Electrolysis of molten ionic compounds forms metal compounds into metals using electricity. Examples include lead (Pb), copper (Cu), and silver (Ag) (Table 4.11). The electrolysis of these compounds when melted produces the metal and non-metal.

**Test yourself**
- Copy and complete the table to show the products of electrolysis of some molten ionic compounds.

<table>
<thead>
<tr>
<th>Ionic Compound</th>
<th>Products of Electrolysis</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium chloride (NaCl)</td>
<td>Sodium metal (Na) and chlorine gas (Cl₂)</td>
</tr>
<tr>
<td>Copper sulfate (CuSO₄)</td>
<td>Copper metal (Cu) and sulfuric acid (H₂SO₄)</td>
</tr>
</tbody>
</table>

**Metal extraction**
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 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 heat energy to melt the metal compounds and
  - The high cost of 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 2073°C. The cost of the thermal energy to melt aluminium oxide for electrolysis is very high. However, the aluminium oxide is mixed with a substance called cryolite, the mixture melts at about 970°C and so the cost of thermal energy to melt this mixture is lower.

The electrolyte for the process is made of cryolite, a form of sodium fluoride (Figure 4.20). Aluminium ions (Al⁺³) are attracted to the negative electrode where they gain electrons and form aluminium metal. As the solution is a liquid, the metal is produced as a liquid and is run off at the bottom. Bicarbonate ions (CO₃⁻²) are attracted to the positive electrode where they lose electrons and form oxygen. The oxygen meets with the graphite anode and as the anode burns to produce carbon dioxide. This means that the anode has to be replaced regularly.

Test yourself

1. Give two reasons why some metals cannot be extracted by heating with carbon.
2. Name two metals other than aluminium that are extracted by electrolysis.
3. Give two reasons why the energy cost of extracting metals by electrolysis is so high.

This question is about the extraction of aluminium by electrolysis.

- Identify the compound from which aluminium is extracted.
- Why is the compound mixed with cryolite?
- Why is it necessary to oxidize the positive electrode?
- What happens at the negative electrode?

Electrolysis of aqueous ionic compounds

When an ionic compound is dissolved in water, the products of electrolysis are often different to those of the compound when it is melted. 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, 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 4.13 shows which ions are discharged at each which 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 electrode used.

Table 4.13 Products from the electrolysis of aqueous solutions of ionic compounds.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Positive electrode products</th>
<th>Negative electrode products</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaCl</td>
<td>Na⁺ and Cl⁻</td>
<td>Na⁺ and OH⁻, O₂, H₂</td>
</tr>
<tr>
<td>CaCl₂</td>
<td>Ca²⁺ and Cl⁻</td>
<td>Ca²⁺ and OH⁻, H₂</td>
</tr>
<tr>
<td>AgNO₃</td>
<td>Ag⁺ and NO₃⁻</td>
<td>Ag⁺ and OH⁻, O₂, H₂</td>
</tr>
<tr>
<td>Pb(NO₃)₂</td>
<td>Pb²⁺ and NO₃⁻</td>
<td>Pb²⁺ and OH⁻, H₂</td>
</tr>
</tbody>
</table>

Test yourself

1. Copy and complete the table to show the products of electrolysis of some aqueous solutions of ionic compounds with inert electrodes.

- What are the products of the electrolysis of aqueous copper(II) nitrate solution using inert electrodes, write a half equation for the process of the oxygen electrode and state whether it is an oxidation or reduction process:
  - at the negative electrode:
  - at the positive electrode:

- For the electrolysis of aqueous copper(II) nitrate solution using inert electrodes, write a half equation for the process of the oxygen electrode and state whether it is an oxidation or reduction process:
  - at the negative electrode:
  - at the positive electrode:
Investigation of what happens when aqueous solutions are electrolysed using inert electrodes

**KEY TEAM**
- Hypothesis: A copper electrode changes to copper sulphate.  
- Observation: The copper electrode becomes shiny and the solution changes to copper sulphate.

<table>
<thead>
<tr>
<th>Electrode</th>
<th>Observation</th>
<th>Conclusion</th>
</tr>
</thead>
<tbody>
<tr>
<td>Copper</td>
<td>Shiny</td>
<td>Copper</td>
</tr>
<tr>
<td>Sulfate</td>
<td>Colour</td>
<td>Copper</td>
</tr>
</tbody>
</table>

**Electrolysis**

**Questions**
1. Why are graphite electrodes used?
2. Why are the aqueous solutions boiled 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 passing products at the cathode to complete the table.
6. Summarise the results, for deciding what is formed at the cathode when aqueous solutions of these compounds are electrolysed using inert electrodes.
7. Write down the identity of the passing 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>pH</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gas</td>
<td>2</td>
<td>3</td>
<td>4</td>
<td>5</td>
<td>7</td>
<td></td>
</tr>
</tbody>
</table>

a) Which solution is the most basic?
b) Which solution is the most acidic?
c) Which solution is neutral?
d) Complete the following word equations, or write no reaction if no reaction would take place.

- Iron + oxygen
- Zinc + sulphuric acid
- Magnesium oxide + hydrochloric acid
- Sodium carbonate + nitric acid
- Potassium hydroxide + sulphuric acid
- Grill + magnesium nitrate
- Iron + copper sulphate
- Zinc + iron nitrate
- Metals are extracted from compounds found in ore, which method is used for each of the following metals?
  a) Magnesium  
  b) Electrolysis  
  c) Copper  
  d) Zinc

e) Grill

2. Balance the following equations:
   a) CaO + H₂O → Ca(OH)₂
   b) Na₂CO₃ + H₂SO₄ → Na₂SO₄ + H₂O
   c) Mg(NO₃)₂ + Fe → MgO + Fe(NO₃)₂

3. 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 ion equation for this reaction.
   d) What type of reaction is this?
   e) The solid magnesium sulfate is heated in an air oven, which is used as a source of combustion. Copy the text to find out the purpose of the magnesium sulfate.

4. Describe how the crystalisation of magnesium sulphate could be carried out.

   a) Write a word equation for this reaction.
Chapter nine: Equation

7. Copy and complete the following table to show the products of electrolysis of some ionic compounds:

<table>
<thead>
<tr>
<th>Ionic Compound</th>
<th>Product at the Negative Electrode</th>
<th>Product at the Positive Electrode</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{H}_2 \text{O} )</td>
<td>( \text{H}_2 )</td>
<td>( \text{O}_2 )</td>
</tr>
<tr>
<td>( \text{AgNO}_3 )</td>
<td>( \text{Ag} )</td>
<td>( \text{NO}_3^- )</td>
</tr>
<tr>
<td>( \text{CuSO}_4 )</td>
<td>( \text{Cu} )</td>
<td>( \text{SO}_4^{2-} )</td>
</tr>
<tr>
<td>( \text{NaCl} )</td>
<td>( \text{H}_2 \text{Cl}_2 )</td>
<td>( \text{Na}^+ )</td>
</tr>
</tbody>
</table>

10. In metal production, when zinc oxide reacts with sulfuric acid, a displacement reaction occurs: \( \text{ZnO} + \text{H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 + \text{H}_2\text{O} \). Explain why aluminum dissolves in acid. (3 marks)

11. a) In the electrolysis of molten sodium chloride, identify the products at the electrodes. (3 marks)
   b) Write a half-equation for the process at the negative electrode and state whether it is an oxidation or a reduction process. (3 marks)
   c) Write a half-equation for the process at the positive electrode and state whether it is an oxidation or a reduction process. (3 marks)
   d) Identify why the anode electrode is made of platinum. (3 marks)

12. Write a balanced equation for the reaction between sulfuric acid and sodium carbonate: \( \text{H}_2\text{SO}_4 + \text{Na}_2\text{CO}_3 \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O} + \text{CO}_2 \). (3 marks)

13. Write a balanced equation for the reaction between copper(II) chloride and sodium carbonate: \( \text{CuCl}_2 + \text{Na}_2\text{CO}_3 \rightarrow \text{CuCO}_3 + \text{NaCl} \). (3 marks)

Practice questions

1. Which one of the following substances does not react with sulfuric acid? (3 marks)
   a) \( \text{H}_2\text{O} \)
   b) \( \text{Zn} \)
   c) \( \text{Cu} \)
   d) \( \text{H}_2\text{SO}_4 \)

2. The acidity of an electrolyte is due to the movement of particles through the substance. What are these particles that move in the electrolyte? (3 marks)
   a) atoms
   b) electrons
   c) ions
   d) protons

3. Complete the following table to give the names and formulas of the ions present in all acids and alcohols: (3 marks)

<table>
<thead>
<tr>
<th>Acid/Alcohol</th>
<th>Ion Name</th>
<th>Ion Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{H}_2\text{SO}_4 )</td>
<td>( \text{H}^+ )</td>
<td>( \text{SO}_4^{2-} )</td>
</tr>
<tr>
<td>( \text{CH}_3\text{OH} )</td>
<td>( \text{CH}_3\text{OH}^- )</td>
<td>( \text{H}^+ )</td>
</tr>
</tbody>
</table>

4. In the electrolysis of copper(II) chloride, what is the product obtained at the anode? (3 marks)
   a) copper
   b) chlorine
   c) hydrogen
   d) oxygen

5. Neutralize excess when an acid and an alkali react to form a salt and water. (3 marks)

6. Copy and complete the table below to give the names and formulas of the ions present in all acids and alcohols: (3 marks)

<table>
<thead>
<tr>
<th>Acid/Alcohol</th>
<th>Ion Name</th>
<th>Ion Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{H}_2\text{SO}_4 )</td>
<td>( \text{H}^+ )</td>
<td>( \text{SO}_4^{2-} )</td>
</tr>
<tr>
<td>( \text{CH}_3\text{OH} )</td>
<td>( \text{CH}_3\text{OH}^- )</td>
<td>( \text{H}^+ )</td>
</tr>
</tbody>
</table>

Use present in all acids: use present in alkalis: sodium carbonate, sodium hydroxide, sulfuric acid.
9 Hydrochloric acid reacts with sodium hydroxide and with sodium carbonate. Compare and contrast the reaction of hydrochloric acid with sodium hydroxide with the reaction of hydrochloric acid with sodium carbonate. In your answer, you must include the names of all products for each reaction and the observations for each reaction. [4 marks]

10 Aluminium is the most abundant metal in the Earth’s crust. Aluminium was 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? [1 mark]
(b) Name the ore from which aluminium is extracted. [1 mark]
(c) The electrolysis of the purified ore is carried out in the Hall-Héroult cell. The diagram below shows the cell used.

(i) Name X and Z. [2 marks]
(ii) Y is the electrolyte. Name the substance in the electrolyte. [2 marks]
(iii) Why is the electrolyte cell kept at about 900°C? [2 marks]
(iv) Note 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]
(vii) Give a reason in terms of electrons why the extraction of aluminium in this process is a reduction reaction. [2 mark]

11. Some substances, for example metals like beryllium and actinum, electron thieves, are described as electrophilic. Other substances for example copper metal, are conductors. 

An experiment, to investigate the electrolysis of the electrolyte medium (lead beryllium), was set up as shown in the diagram below.

(a) Some pieces of apparatus in the diagram are labelled A-E. State the correct name for each piece of apparatus. [4 marks]
(b) Name one piece of apparatus which should be connected to the circuit to show that an electric current is flowing through the electrolyte [1 mark]
(c) Copy and complete the table to state the names of the products, and the half equations for the electrolytes.

<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>
<tr>
<td>C</td>
<td></td>
<td></td>
</tr>
<tr>
<td>D</td>
<td></td>
<td></td>
</tr>
<tr>
<td>E</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

- Why does this electrolyte need to be carried out in a tube? [2 marks]
- Based on the difference in conductivity of electrolytes, copper metal and metallic lead beryllium. [2 marks]
- Give one product of each electrode when aqueous sodium chloride is electrolysed. [2 marks]

**Working scientifically: Measurements and uncertainties**

When we 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 commonly in chemistry include:

- **Measuring cylinder** (Figure 4.27)
- **pipette** (Figure 4.28)
- **buoyette**
- **thermometer**
- **balance**

- **Measuring cylinders:** pipettes and burettes are used to measure not volumes of liquids, but small volumes of gases, which are used in titrations (see Chapter 5). 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 4.27).

**FYE 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 52, 77, 76, 80 and 82 cm³.

The mean value is found after rejecting any anomalous results. The cm³ is anomalous here as it is significantly different from all the others.

Mean volume = \( \frac{52 + 77 + 76 + 80 + 82}{5} = 76 \text{ cm}^3 \)

The mean is quoted to the nearest unit as all the values are measured to the nearest unit. The accuracy is ± 1 cm³ as the highest and lowest values are within ± 1 cm³ of the mean.

Questions

1. Record the volume of liquid in each measuring cylinder A-D. All scales are shown in cm³.

2. Record the volume of liquid in each of the containers E to G. All scales are shown in cm³.

3. Record the temperature shown on each thermometer in K. All scales are shown in °C.

4. In each pair of figures, find the difference in mass between the first reading and the second reading. All scales are shown in grams.

5. Find the mean value and its uncertainty for the boiling point of a substance which was measured several times. Values measured were 102, 103, 105 and 102°C.

6. Find the mean value and its uncertainty for the mass of a substance which was measured several times. Values measured were 25.6, 25.8, 25.7 and 25.5 g.

7. Find the mean value and its uncertainty for the volume of a gas measured in a reaction. Values measured were 52, 77, 76, 80 and 82 cm³.

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 1a, 1b, 1c, 2a and 2b and is called energy changes. It covers exothermic and endothermic reactions, as well as chemical cells and fuel cells.
Exothermic and endothermic reactions

Previously you could have learned:

- 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.
- Reactions are substances that are good sources of energy.

Test yourself on prior knowledge:

- Chemical energy and thermal energy are two forms of energy.
- Write down three forms of energy besides these two.
- What is the use of conservation of energy?
- What is a heat?
- Name three common fuels.

Exothermic reactions

In exothermic reactions, thermal energy is transferred from the chemicals to the surroundings. As there is less thermal energy, the temperature decreases and it feels cooler. Most chemical reactions are exothermic.

- In some exothermic reactions, only a small amount of thermal energy is transferred and the temperature may only fall 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 5.1).

Table 5.1: Examples of exothermic reactions

<table>
<thead>
<tr>
<th>Reaction</th>
</tr>
</thead>
</table>
| Condensation, e.g. | Chemical energy is transferred from the chemicals to the surroundings. As there is less thermal energy, the temperature decreases and it feels cooler.
| Combustion reactions e.g. | Chemical energy is transferred from the chemicals to the surroundings. As there is less thermal energy, the temperature decreases and it feels cooler.
| Reaction of acids with metal | Chemical energy is transferred from the chemicals to the surroundings. As there is less thermal energy, the temperature decreases and it feels cooler.
| Decomposition reactions | Chemical energy is transferred from the chemicals to the surroundings. As there is less thermal energy, the temperature decreases and it feels cooler.

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 it feels cooler. Endothermic reactions are less common than exothermic reactions.

Applications of endothermic reactions

Some sports injury pads act as a cold pack to put on an injury to prevent swelling (Figure 5.5). 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.
Test yourself

1. Which of the following reactions endothermic or exothermic?
   a) The temperature increased at 2°C and finished at 4°C.
   b) The temperature started at 1°C and finished at 2°C.
   c) The temperature was constant at 2°C. 0°C. 0°C.
   d) The temperature decreased at 2°C and finished at 1°C.

2. Which of the following reactions endothermic or exothermic?
   a) heating a sample of calcium carbonate.
   b) Reaction of sugar and water to make sugar water.
   c) Reaction of calcium oxide and water to make calcium hydroxide.
   d) Reaction of iron and oxygen to make iron oxide.

3. Show you can... (Table)

4. 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.

   ![Figure 5.1 Reaction profiles for exothermic and endothermic reactions.](image)

5. 5. Energy changes

   Investigating the variables that affect temperature change in reacting solutions - the temperature change in a neutralisation reaction

   In this experiment the following method was followed:
   1. A solution of sodium hydroxide was prepared in a plastic beaker.
   2. A burette was filled with 5 drops of hydrochloric acid.
   3. The temperature of the sodium hydroxide was measured and recorded.
   4. The burette was then transferred to the beaker.
   5. The mixture was then stirred and the temperature was measured and recorded.
   6. The experiment was repeated.

   The results are shown in Table 5.1.

<table>
<thead>
<tr>
<th>Volume of acid added ml</th>
<th>Temperature change °C</th>
</tr>
</thead>
<tbody>
<tr>
<td>5</td>
<td>2.4</td>
</tr>
<tr>
<td>10</td>
<td>3.2</td>
</tr>
<tr>
<td>15</td>
<td>4.5</td>
</tr>
<tr>
<td>20</td>
<td>5.5</td>
</tr>
<tr>
<td>25</td>
<td>5.8</td>
</tr>
<tr>
<td>30</td>
<td>6.0</td>
</tr>
<tr>
<td>35</td>
<td>6.2</td>
</tr>
<tr>
<td>40</td>
<td>6.3</td>
</tr>
</tbody>
</table>

6. Questions

   a) In this experiment identify the independent variable.
   b) Dependent variable.
   c) Key control variables.
   d) Why was the experiment repeated?
   e) Why was a polythene cup used to hold the mixture?
   f) Why should the burette be allowed to cool before each addition of acid?
   g) Why should a graph of average temperature Vs volume of acid be plotted?

   Describe the trend shown by the results plotted.

   7. In this experiment the highest temperature is reached when one complete neutralisation has occurred. Imagine you would experimentally determine that the temperature of a single solution containing an excess of acid was much higher than the temperature of the neutralisation that occurred. A different experiment was carried out by using two solutions of different acids which were allowed to react in the same way. The highest temperature reached was recorded and presented in Table 5.4.

   8. Describe the trend shown by the results plotted.

   9. In this experiment identify the independent variable.
   a) Dependent variable.
   b) Key control variables.
   c) State two conclusions you can draw from the results.

   ![Figure 5.2 Temperature change in reacting solutions.](image)
**Bond Energies**

- Breaking a chemical bond takes energy. For example, 436.6 kJ of energy is needed to break one mole of H–H covalent bonds. Due to the law of conservation of energy, 436.6 kJ of energy must be released when making one mole of H–H covalent bonds (Figure 5.5).

- Dying 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 bonds in the reactants and the energy released when bonds are made in the products.

---

**Example**

Find the energy change in the following reaction using the bond energies given.

<table>
<thead>
<tr>
<th>Bond energies: C≡C = 3.72 kJ/mol, O≡O = 595.9 kJ/mol, O=O = 154.8 kJ/mol</th>
</tr>
</thead>
</table>

**A**

- Energy change = energy needed to break bonds – energy released making new bonds
- Change in energy = -3.72 kJ/mol

---

**Example**

Find the energy change in the following reaction using the bond energies given.

<table>
<thead>
<tr>
<th>Bond energies: C≡C = 3.72 kJ/mol, O≡O = 595.9 kJ/mol, O=O = 154.8 kJ/mol</th>
</tr>
</thead>
</table>

**B**

- Energy change = energy needed to break bonds – energy released making new bonds
- Change in energy = -3.72 kJ/mol

---

**Example**

Find the energy change in the following reaction using the bond energies given.

<table>
<thead>
<tr>
<th>Bond energies: C≡C = 3.72 kJ/mol, O≡O = 595.9 kJ/mol, O=O = 154.8 kJ/mol</th>
</tr>
</thead>
</table>

**C**

- Energy change = energy needed to break bonds – energy released making new bonds
- Change in energy = -3.72 kJ/mol

---

**Example**

Find the energy change in the following reaction using the bond energies given.

<table>
<thead>
<tr>
<th>Bond energies: C≡C = 3.72 kJ/mol, O≡O = 595.9 kJ/mol, O=O = 154.8 kJ/mol</th>
</tr>
</thead>
</table>

**D**

- Energy change = energy needed to break bonds – energy released making new bonds
- Change in energy = -3.72 kJ/mol

---

**Energy changes**

- Endothermic reaction: energy is absorbed when bonds are broken.
- Exothermic reaction: energy is released when bonds are formed.

---

**Test yourself**

- Look at the following reaction profiles.
- Reaction 1:
  - Progress of reaction
  - Energy of activation
  - Reaction 2:
  - Progress of reaction
  - Energy of activation
  - Reaction 3:
  - Progress of reaction
  - Energy of activation

- a) Which reaction takes the most activation energy for the conversion of reactants to products?
- b) Which reaction is endothermic?
- c) Which reaction is exothermic?
- d) Which reaction takes the least activation energy for the conversion of reactants to products?

**Answer**

- Reaction 3 takes the most activation energy.
- Reaction 2 is endothermic.
- Reaction 1 is exothermic.
- Reaction 2 takes the least activation energy.
Chemical cells and fuel cells

Chemical cells

We use cells and batteries every day. They contain chemicals which want to supply electricity.

A very simple chemical cell can be made using a lemon and two pieces of different metals (Figure 5.7). The lemon juice acts as an electrolyte. An electrolyte is a liquid that conducts electricity.

Salt solution is another example of a good electrolyte. In most electrolytes, ions (not electrons) carry the electric charge through the liquid. For example, in sodium chloride solution, it is the sodium ions and chloride ions that move through the solution carrying the electric charge.

Chemical cells can be made by placing two different metals in a beaker containing sodium chloride solution as electrolyte (Figure 5.8). There is a potential difference (voltage) between the two metals, so electrons flow through the atoms from the more reactive metal to the less reactive metal, while ions from sodium chloride carry the electric charge through the solution.

There are several factors that affect the voltage produced by the cell. These include the identity of the metals and the electrolyte.

1. The metals: The identity of the metals used as electrodes in the cell makes a significant difference. The greater the difference in reactivity between the two metals, the greater the voltage.
2. The electrolyte: Changing the electrolyte and/or the concentration of the electrolyte can affect the voltage.

Test yourself

1. A simple chemical cell can be made from two different metals placed in an electrolyte.
   a. What is an electrolyte?
   b. Name a good electrolyte.
   c. Define the terms voltage and electric charge.
   d. In a simple chemical cell, how is the electric charge carried through the electrolyte?
   e. What happens when the two electrodes are connected?
   f. State two factors that affect the potential difference (voltage) of a cell.

2. Why would a cell made from two pieces of copper joined by a wire and placed in an electrolyte not produce a potential difference (voltage)?

Tip

Remember that energy is released in a redox reaction but is released when bonds are made.

Show you can...

The energy change is –1164 kJ in the following reaction:

\[ \text{C} + \text{O}_2 \rightarrow \text{CO}_2 \]

Bond energies: \( \text{H-H} = 436 \text{ kJ/mol} \)
\( \text{C} = 348 \text{ kJ/mol} \)
\( \text{O}_2 = 496 \text{ kJ/mol} \)
\( \text{C} + \text{O}_2 = 844 \text{ kJ/mol} \)

Calculate the bond energy of \( \text{C} + \text{O}_2 \)
### Batteries

The voltage produced by many chemical cells is typically 1.5 volts. In order to increase this voltage, several cells can be connected together. A battery is a two or more cells connected together in series to provide a greater voltage. A 4 volt battery is made from four 1.5 volt cells connected together (Figure 5.3).

As current is drawn from a cell, the chemicals inside the cell are used up. If non-rechargeable cells and batteries, once one or more of the chemicals has been used up the battery is "flat" and no longer works. Alkaline cells and lithium button cells are non-rechargeable (Figure 5.5.5).

In rechargeable cells and batteries, the chemical reactions can be reversed to use these chemicals. This means that they can be used over and over again. The cells and batteries in mobile phones, tablets and laptops are rechargeable (Figure 5.11). The chemical reactions inside the cell can be reversed by passing an electric current through the cell.

### Evaluating the use of cells and batteries

Cells and batteries are very useful, but there are some problems with them. Table 5.6 evaluates the use of cells and batteries.

---

### Test yourself

1. What is a battery?
2. What is the reason for making batteries from cells?
3. Why is a battery and its components used up?
4. What does it mean to use a battery in a circuit?
5. How can you use a battery to do useful work?
6. One of the most common cells used is the alkaline AA 1.5 volt cell. Is this rechargeable or non-rechargeable?
7. Why is it important that used cells and batteries are not thrown away in household rubbish?
8. **Fuel cells**

   Non-rechargeable cells and batteries contain a limited amount of chemicals that are used up. However, in a fuel cell there is a supply of the chemicals into the cell replacing those that are used up and so the cell keeps working. The chemicals are usually oxygen (air which contains oxygen) plus a fuel which is usually hydrogen. In a hydrogen fuel cell, an electrochemical reaction takes place between hydrogen and oxygen producing a potential difference. The hydrogen is oxidised and the only product of the reaction is water. This means that hydrogen fuel cells do not produce any polluting gases.

   The overall reaction in a hydrogen fuel cell is:

   \[
   \text{hydrogen (H}_2\text{)} + \text{oxygen (O}_2\text{)} \rightarrow \text{water (H}_2\text{O)}
   \]

   This reaction is very efficient and the electrical energy supplied can be used to power electrical devices or even power motors in vehicles such as cars and buses (Figures 5.12 and 5.13).
What happens inside a fuel cell?

- At one of the electrodes, hydrogen gas loses electrons and forms hydrogen ions. The half equation for this is: \(2H_2 + 4e^- \rightarrow 2H^+\).
- The hydrogen ions move to the other electrode through the electrolyte.
- The electrons move to the other electrode through a wire. This current can be used to power a connected device.
- The hydrogen ions and electrons react with the oxygen gas at the other electrode making water. The half equation for this is: \(2H^+ + O_2 + 2e^- \rightarrow H_2O\) (Figure 5.5).

Evaluating the use of fuel cells

Fuel cells are very useful but there are some problems with their use. Table 5.7 evaluates the use of fuel cells compared to rechargeable cells and batteries.

Test yourself

1. What is the key difference between a fuel cell and other types of cell?
2. Write an overall equation for the reaction that takes place in a hydrogen fuel cell.
3. Why are hydrogen fuel cells categorised as exothermic? Explain why this term is particular as it is not actually true.
4. What is the main reason that fuel cells are not used more than they are?

Show you care

The table shows the percentage error at the anode and the cathode for a hydrogen fuel cell.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Anode</th>
<th>Cathode</th>
<th>Overall</th>
</tr>
</thead>
<tbody>
<tr>
<td>(2H_2 \rightarrow 4H^+ + 4e^-)</td>
<td>(4e^-)</td>
<td>(2H_2 + 4e^- \rightarrow 2H^+)</td>
<td></td>
</tr>
</tbody>
</table>

Chapter review questions

1. The following reactions take place in solution in a beaker. The temperature before and after the reactions were measured and recorded in each case. Decide whether each reaction is exothermic or endothermic.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Initial Temperature</th>
<th>Final Temperature</th>
<th>Change</th>
</tr>
</thead>
<tbody>
<tr>
<td>Reaction 1</td>
<td>10°C</td>
<td>15°C</td>
<td>5°C</td>
</tr>
<tr>
<td>Reaction 2</td>
<td>20°C</td>
<td>25°C</td>
<td>5°C</td>
</tr>
<tr>
<td>Reaction 3</td>
<td>30°C</td>
<td>35°C</td>
<td>5°C</td>
</tr>
</tbody>
</table>

2. Copy and complete the spaces in the following sentences.

In the exothermic reaction, thermal energy is transferred from the chemical to the surroundings and so the temperature __________.

In the endothermic reaction, thermal energy is absorbed from the surroundings and so the temperature __________.

3. Copy and complete the spaces in the following sentence: Chemical reactions can only take place when particles __________ with each other and have enough energy. The maximum energy particles need to react is called the __________.

4. The reaction profile for a reaction is shown.

5. Decide whether each of the following reactions is likely to be exothermic or endothermic.

- heat-up magnesium
- decomposition of silver oxide
- reaction inside a copper-iron cell pack
- reaction inside a self-heating food can
- neutralisation of sulfuric acid by sodium hydroxide
- neutralisation of sulfuric acid by sulfuric hydrogen peroxides

6. The table shows the potential difference (voltage) when chemical cells were set up by allowing two strips of metal into a beaker of salt solution. In each case the copper was connected to the positive terminal of a voltmeter.

<table>
<thead>
<tr>
<th>Electrode</th>
<th>Potential (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Copper</td>
<td>Zinc</td>
</tr>
<tr>
<td>Magnesium</td>
<td>-0.75</td>
</tr>
<tr>
<td>Nickel</td>
<td>+0.30</td>
</tr>
</tbody>
</table>

Chapter 5
Chemical cells and fuel cells

Test yourself

1. What is the key difference between a fuel cell and other types of cell?
   a. It produces chemical products from a hydrogen fuel cell.
   b. It produces chemical products from a hydrogen fuel cell.
   c. It produces chemical products from a hydrogen fuel cell.
   d. It produces chemical products from a hydrogen fuel cell.

2. Write an overall equation for the reaction that takes place in a hydrogen fuel cell.
   a. $2H_2 + O_2 \rightarrow 2H_2O$
   b. $2H_2 + O_2 \rightarrow 2H_2O$
   c. $2H_2 + O_2 \rightarrow 2H_2O$
   d. $2H_2 + O_2 \rightarrow 2H_2O$

3. What is the term that describes the process of converting energy from one form to another in a fuel cell?
   a. Electrolysis
   b. Electrocatalysis
   c. Electrocatalysis
   d. Electrocatalysis

4. What is the main reason that fuel cells are not used more than they are?
   a. They are too expensive
   b. They are too inefficient
   c. They are too complex
   d. They are too complex

5. What is the key difference between a fuel cell and other types of cell?
   a. It produces chemical products from a hydrogen fuel cell.
   b. It produces chemical products from a hydrogen fuel cell.
   c. It produces chemical products from a hydrogen fuel cell.
   d. It produces chemical products from a hydrogen fuel cell.

Chapter review questions

1. The following reactions take place in a fuel cell. The temperature before and after the reaction were measured to be 25°C. Decide whether each reaction is endothermic or exothermic:
   a. $2H_2 + O_2 \rightarrow 2H_2O$
   b. $2H_2 + O_2 \rightarrow 2H_2O$
   c. $2H_2 + O_2 \rightarrow 2H_2O$
   d. $2H_2 + O_2 \rightarrow 2H_2O$

2. Copy and complete the spaces in the following sentences:
   a. In exothermic reaction, heat energy is transferred from the chemical to the surroundings.
   b. In exothermic reaction, heat energy is transferred from the chemical to the surroundings.

3. Copy and complete the spaces in the following sentences:
   a. In exothermic reaction, heat energy is transferred from the chemical to the surroundings.
   b. In exothermic reaction, heat energy is transferred from the chemical to the surroundings.

4. Copy and complete the spaces in the following sentences:
   a. In exothermic reaction, heat energy is transferred from the chemical to the surroundings.
   b. In exothermic reaction, heat energy is transferred from the chemical to the surroundings.

5. Copy and complete the spaces in the following sentences:
   a. In exothermic reaction, heat energy is transferred from the chemical to the surroundings.
   b. In exothermic reaction, heat energy is transferred from the chemical to the surroundings.

6. Decide whether each of the following reactions is likely to be exothermic or endothermic:
   a. Burning magnesium
   b. Decomposition of silver oxide
   c. Reaction inside a spark from a spark plug
   d. Reaction inside a self-heating food can
   e. Neutralization of sulfuric acid by sodium hydroxide
   f. Neutralization of sulfuric acid by sodium hydroxide
   g. Neutralization of sulfuric acid by sodium hydroxide
   h. Neutralization of sulfuric acid by sodium hydroxide

7. The table shows the potential differences (volts) when chemical cells were set up. The two components of each cell were connected to the positive terminals of a voltmeter.
   a. $Zn$ and $Cu$ at 1.10
d. $Ag$ and $Cu$ at 0.79

8. The table shows the potential differences (volts) when chemical cells were set up. The two components of each cell were connected to the positive terminals of a voltmeter.
   a. $Zn$ and $Cu$ at 1.10
   b. $Ag$ and $Cu$ at 0.79
   c. $Cu$ and $Ag$ at 0.58

9. The table shows the potential differences (volts) when chemical cells were set up. The two components of each cell were connected to the positive terminals of a voltmeter.
   a. $Zn$ and $Cu$ at 1.10
   b. $Ag$ and $Cu$ at 0.79
   c. $Cu$ and $Ag$ at 0.58

10. The table shows the potential differences (volts) when chemical cells were set up. The two components of each cell were connected to the positive terminals of a voltmeter.
    a. $Zn$ and $Cu$ at 1.10
    b. $Ag$ and $Cu$ at 0.79
    c. $Cu$ and $Ag$ at 0.58
Practice questions

1. What is an electrolyte? [1 mark]
2. What is the electrolyte in this experiment? [1 mark]
3. How is the silver metal in the table in order of reactivity, clearly showing where it is the most reactive? [1 mark]
4. What would be the potential difference if two pieces of copper metal were used as the electrodes? [1 mark]
5. What would be the potential difference if zinc and magnesium were used as the electrodes? [1 mark]
6. Cells and batteries are very useful portable sources of electricity. Some cells and batteries are rechargeable. A what is a battery? [1 mark]
7. Why do cells and batteries go flat? [1 mark]
8. What happens inside rechargeable cells and batteries when they are recharged? [1 mark]
9. Give one problem with the use of cells and batteries. [1 mark]

8. a) Calculate the energy change for the following reaction using these bond energies.
   \[
   \begin{align*}
   N\equiv N & \rightarrow 2N\equiv \Delta H = -157 \text{ kJ/mol} \\
   O\equiv O & \rightarrow 2O\equiv \Delta H = -143 \text{ kJ/mol} \\
   N\equiv O & \rightarrow NO \Delta H = \text{? kJ/mol}
   \end{align*}
   
   b) In this reaction endothermic or exothermic? Explain your answer by assigning bond energies. [1 mark]

9. Calculate the energy change for the following reaction using these bond energies.
   \[
   \begin{align*}
   H\rightarrow H & \rightarrow 2H\equiv \Delta H = 436 \text{ kJ/mol} \\
   O\rightarrow O & \rightarrow 2O\equiv \Delta H = 496 \text{ kJ/mol} \\
   H\rightarrow O & \rightarrow \text{H}_2O \Delta H = \text{? kJ/mol}
   \end{align*}
   
   a) What is the reaction type? [1 mark]
   b) 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. [1 mark]
   - combustion
   - disproportionation
   - neutralization
   - reduction
   - oxidation
   c) 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. [1 mark]
   - combustion
   - disproportionation
   - neutralization
   - reduction
   - oxidation
   d) Explain the difference between an exothermic reaction and an endothermic reaction. [1 mark]
   e) For each of these reactions A to E, state which one is an exothermic reaction. [1 mark]
   f) For each of these reactions A to E, state which one is an endothermic reaction. [1 mark]
   g) Describe how you would experimentally prove that the reaction between magnesium and hydrochloric acid is an exothermic reaction. [1 mark]
   h) Photosynthesis is an exothermic process used by plants to convert light energy into chemical energy stored in sugar molecules. \[
   \text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6 + \text{O}_2
   \]
   i) What is meant by the term exothermic? [1 mark]
   j) Describe what is meant by the term activation energy. [1 mark]

10. a) 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. [3 marks]
   b) Explain in terms of bond making and breaking why this reaction is exothermic. [2 marks]
   c) Sodium oxide (Na$_2$O) is present in car brakes. During a car crash the sodium rapidly reacts with oxygen gas from the reaction shown below.
   \[
   2\text{Na}_2\text{O}(s) + \text{O}_2(g) \rightarrow 4\text{Na}_2\text{O}(s)
   \]
   The energy change for this reaction is positive. [1 mark]
   d) What is the term given to reactions which have a positive energy change? [1 mark]
   e) Draw a labelled reaction profile for the reaction. [1 mark]
   f) Calculate the mass of sodium oxide needed to produce 1224kPa of oxygen gas at 30°C and one atmosphere pressure. (Carried from Chapter 13.4.1, remember to refer to the table for heat of combustion at 30°C and one atmosphere pressure). [2 marks]
   g) Draw the diagram shown for the reaction with a catalyst and without a catalyst. [1 mark]

11. a) 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. [3 marks]
   b) Explain in terms of bond making and breaking why this reaction is exothermic. [2 marks]
   c) The cop of the diagram shows the activation energy of the catalysed reaction at A and the activation energy for the uncatalysed reaction at B. [1 mark]
   d) From the information shown in the graph, state the effect of a catalyst on the activation energy of a reaction. [1 mark]
Hydrogen reacts with fluorine to form hydrogen fluoride:
\[ H_2 + F_2 \rightarrow 2HF \]

Use the bond energies in the table to calculate the energy change for this reaction. (4 marks)

<table>
<thead>
<tr>
<th>Bond</th>
<th>Bond energy in kJ/mol</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>

Bond in oxygen:
\[ \rightarrow \quad H-F \rightarrow \quad H=F \rightarrow \quad H^{+}+\text{F}^{-} \]

See the bond energies below to calculate the energy change for this reaction. (3 marks)

Bond energies:
- H-H: 436 kJ/mol
- O-O: 158 kJ/mol
- H-O: 568 kJ/mol

Hydrogen is used as a liquid propellant to launch rockets. (2 marks)

a) Explain using a balanced chemical equation why hydrogen is a clean fuel. (2 marks)

b) Hydrogen is also used in fuel cells, which are taken into space to supply power. Explain how a hydrogen fuel cell generates electricity. (3 marks)

Working scientifically: Identifying variables when planning experiments

When carrying out an experiment, different variables are used. Variables are the things that can change. When you plan experiments, you should 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** is a variable that can take any value between two extremes. For example, the temperature or the growth of a plant.
- A **discrete variable** is a variable that can only take certain values. For example, the number of people or the number of goldfish in a pond.
- A **categorical variable** is one which is best described by words. Variables such as the type of acid or the type of metal are categorical variables.

### Questions

1. Identify the following as categorical or continuous variables.

   - Temperature
   - Mass
   - Volume
   - Time
   - Molarity
   - pH
   - Type of acid
   - Number of goldfish

2. 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. the one which you intentionally change during an experiment).

   - The dependent variable is the variable which may change as a result of changing the independent variable. It is the one which is measured for each value of the independent variable.

   A **causal relationship** is one which may involve addition to the independent variable, affect the outcome of the experiment. Causal variables must be kept constant during an experiment to make it a fair test.
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 carbonate and hydrochloric acid.

**Independent variable**

- Temperature

**Dependent variable**

- The rate at which magnesium carbonate reacts

**Controlled variables**

- The mass of magnesium carbonate
- The volume of hydrochloric acid
- The concentration of hydrochloric acid

### Questions

1. For the following experiments identify the
   - Independent variable
   - Dependent variable
   - Controlled variables

   a. Some hydrogen 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 displace gas was recorded. The experiment was repeated using different masses of copper carbonate.

   c. 2g of magnesium was added to copper sulphate solution and the highest temperature reached was recorded. The experiment was repeated using different masses of magnesium.

   d. In an experiment to find the effect of stirring on speed of dissolving, the time taken to dissolve some copper sulphate in water was measured. This was repeated using 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 ice and water was recorded before and after some solid ammonium nitrate was added. The experiment was repeated using solid ammonium nitrate in different masses.

---

6 The rate and extent of chemical change

The explosion of dynamite is a very fast reaction. The rusting of steel is a very slow reaction. Some reactions completely react, but in others the products can turn back into the reactants. Being able to control the speed of a reaction and reducing the amount of products turning back into reactants are very important and are studied in this chapter.

This chapter covers specification points 6.1.1 to 6.1.3 and is called 'The rate and extent of chemical change.' It covers the rates of reactions, reversible reactions and dynamic equilibrium.
Previously you could have learned:

- Some reactions are fast and some reactions are slow.
- Some chemical changes are irreversible.
- Exothermic reactions release thermal energy to the surroundings and as the temperature increases.
- Endothermic reactions take in thermal energy from the surroundings and as the temperature decreases.

Test yourself on prior knowledge:

1. Give two examples of chemical reactions that are very fast.
2. Describe endothermic and exothermic reactions.
3. Explain why the temperature falls in an exothermic reaction.
4. Explain why the temperature falls in an endothermic reaction.

Rate of reaction

Measuring the rate of reaction

Some chemical reactions are fast, while others are slow. Reactions that are fast have a high rate of reaction. Reactions that are slow have a low rate of reaction.

Measuring the mean rate of reaction

The mean rate of a reaction can be found by measuring the quantity of a reactant used or a product formed over the time taken.

mean rate of reaction = \frac{quantity of reactant used}{time} or \frac{mean rate of reaction = \frac{quantity of product formed}{time}}

The time is typically measured in seconds.

The quantity of a chemical could be measured as:

- mass in grams (g) or
- volume in cubic centimeters (cm³) or
- amount in moles (mol).

The rate of a reaction changes during a reaction. Most reactions are fast at the beginning, slow down and then eventually stop. As the rate is constantly changing, the rate at one moment is likely to be different to the rate at another. This means that when we calculate the rate of reaction over a period of time, we are actually working out the mean rate of reaction over that time.
Reaction rate graphs

Graphs can be drawn to show how the quantity of reactant used or product formed changes with time. The shapes of these graphs are usually very similar for most reactions. The slope of the line represents the rate of the reaction. The steeper the slope, the faster the reaction.

The relative slope can be judged by simply looking at the line but it can be useful to draw tangents to the curve (Figure 6.1). At any point on a curve, a straight line is drawn with a ruler that just touches the curve and then we judge to be at the same slope as the curve at that point. The steeper the tangent, the faster the rate of reaction.

In an individual reaction, the line is steepened at the start when the reaction at its fastest. The line becomes less steep as the reaction slows down and eventually becomes horizontal when the reaction stops. In reactions that produce gas, it is easy to measure the volume or mass of gas formed over time (Table 6.1) and draw a graph of this type.

The rate of one reaction can be compared to the rate of another using these graphs. The steeper the slope of the line, the greater the rate of reaction. For example, in the graph in Figure 6.2, reaction A is faster than reaction B. We can see that the line for reaction A is steeper, but this is confirmed if we draw a tangent to each curve near the start.

Finding the rate of reaction from the gradient of tangents on reaction rate curves

The rate of reaction at any given moment can be found by measuring the gradient of the tangent to the curve at that point. To calculate the rate of reaction by drawing a tangent to a graph:

1. Select the time at which you want to measure the rate.
2. With a ruler, draw a tangent to the curve at that point (the tangent should have the same slope as the graph line at that point).

Test yourself

The graph shows how the mass of carbon dioxide formed when sodium carbonate reacts with hydrochloric acid varies over time.

1. At what point on the graph is the rate of reaction greatest? Explain your answer.
2. At what point on the graph does the reaction stop? Justify your answer.
3. At the rate of reaction greater at point A or point B? Explain your answer.
4. What is the rate of reaction of reaction A compared to reaction B?
5. Which reaction is the fastest?
6. At which point is the reaction slower?

Hydrogen gas is formed when magnesium reacts with sulfuric acid.

The table shows how the volume of hydrogen increased with time when some magnesium was reacted with sulfuric acid.

<table>
<thead>
<tr>
<th>Time (s)</th>
<th>Volume of Hydrogen (cm³)</th>
</tr>
</thead>
<tbody>
<tr>
<td>10</td>
<td>20</td>
</tr>
<tr>
<td>20</td>
<td>30</td>
</tr>
<tr>
<td>30</td>
<td>40</td>
</tr>
<tr>
<td>40</td>
<td>50</td>
</tr>
<tr>
<td>50</td>
<td>60</td>
</tr>
</tbody>
</table>

1. Find the mean rate of reaction over 10 seconds. Calculate the mean rate of reaction during the time intervals at 20 seconds.
2. Calculate the rate of reaction at each time interval and draw a graph of the results.
3. Calculate the rate of reaction at 20 seconds by drawing a tangent to the curve.
4. Write an equation for this reaction.
Collision theory

For a chemical reaction to take place particles of the reactants must collide with enough energy to react. The minimum amount of energy particles need to react is called the activation energy.

Successful collisions (i.e. ones which result in a reaction) take place when reactant particles collide with enough energy to react (Figure 4.4). Unsuccessful collisions, even which do not result in a reaction, take place when reactant particles collide but do not have enough energy to react.

The rate of a reaction depends on the frequency of successful collisions.

Factors affecting the rate of reactions

There are several factors that affect the rate of chemical reactions.

Temperature

The higher the temperature, the faster a reaction. This is because the particles have more energy (they are more energetic) and so more of the particles have enough energy to react when they collide. The particles are also moving faster and so collide more frequently. This means that there are more frequent successful collisions.

An everyday example of the effect of temperature on chemical reactions is the rate of freezing (Figure 6.5). When food goes off, chemical reactions are taking place. We can slow down the rate at which these reactions take place and so slow down the rate at which food goes off by cooling food down and putting it in a fridge.

Concentration of reactants in solution

Many chemical reactions involve reactants that are dissolved in solutions. The concentration of a solution is a measure of how much solute is dissolved. The higher the concentration, the more particles of solute that are dissolved (Figure 6.6).

The higher the concentration of reactants in solution, the greater the rate of reaction. This is because there are more reactant particles in the solution and so there are more frequent successful collisions.

Pressure of reacting gases

Some chemical reactions involve reactants that are gases. The higher the pressure of a gas, the closer the reactant particles are together.

The higher the pressure of reactants that are gases, the greater the rate of reaction. This is because the reactant particles are closer together and so there are more frequent successful collisions (Figure 6.7).

Surface area of solid reactants

Some chemical reactions involve reactants that are solids. The surface area of a solid is increased if it is broken up into more pieces (Figure 6.8). When a solid is made into a powder it has a massive surface area. The greater the surface area of a solid reactant, the greater the rate of reaction. This is because there are more particles on the surface that can react and so there are more frequent successful collisions.

A solid cube with sides of length 2 cm has a total surface area of 24 cm². If it is broken up into eight smaller cubes with sides of length 1 cm then the surface area increases to 64 cm². This means that the surface area has increased while the total volume stays the same (Table 6.2).

The surface area to volume ratio can also be considered to explain this.

The single cube with four sides has a surface area to volume ratio of 3:1. When it is broken up into eight smaller cubes with 1 cm sides, the surface area to volume ratio is greater than 6:1 (Table 6.2).
Test yourself!

1. Hydrogen peroxide decomposes very slowly in water and oxygen. The reaction is much faster in a catalyst such as manganese peroxide is it used. When is a catalyst necessary?
2. How does a catalyst work?
3. This diagram shows the reaction paths for the decomposition of hydrogen peroxide in the absence of catalyst. Which curve represents the activation energy for the reaction that takes place with a catalyst?
4. The rate of reaction can also be increased by adding an enzyme. What are enzymes?

Nitrogen gas reacts with hydrogen gas to make ammonia. The reaction is carried out at 200 atmospheres pressure. Explain why the reaction takes place at higher pressure.

Sodium thiosulphate solution reacts with hydrochloric acid and sodium thiosulphate solution becomes cloudy. The reaction can be carried out in a test tube or in a test tube holder. Describe an experiment to show that the reaction can be reversed by adding more sodium thiosulphate solution.

Another way of measuring the rate of reaction is by measuring the loss of mass of reactants, often by recording the mass of the reacting vessel and its contents over a certain period of time, as shown in the example in the practical activity on the next page.
Investigating the rate of decomposition of hydrogen peroxide solution by measuring the loss in mass

Hydrogen peroxide decomposes in the presence of solid manganese dioxide to produce water and oxygen.

$$\text{H}_2\text{O}_2 \rightarrow \text{H}_2\text{O} + \text{O}_2$$

The apparatus shown was used to investigate the rate of decomposition of hydrogen peroxide solution. 20cm$^3$ of hydrogen peroxide solution was added to 1g of solid manganese dioxide at 20°C.

The following results were obtained.

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Time for 50% decomposition (s)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>3.34</td>
</tr>
<tr>
<td>2</td>
<td>3.54</td>
</tr>
<tr>
<td>3</td>
<td>3.88</td>
</tr>
<tr>
<td>4</td>
<td>4.18</td>
</tr>
</tbody>
</table>

**Questions:**

a) Write a balanced chemical equation for the decomposition of hydrogen peroxide (H$_2$O$_2$).

b) What is the purpose of the cotton wool plug?

c) Why is a graph of mass of oxygen lost against time drawn?

d) Use the graph to state the mass of oxygen lost after 2 minutes.

e) Suggest an alternative way of measuring the rate of this reaction without measuring the mass of oxygen lost.

f) State on the same axes the graph you would expect to obtain if the experiment was repeated at 30°C using 1g of manganese dioxide with 20cm$^3$ of this hydrogen peroxide solution instead with 20cm$^3$ of water.

2 At the end of this experiment the manganese dioxide can be recovered,

a) Draw a labelled diagram of the assembly of apparatus which could be used to recover the manganese dioxide at the end of the experiment.

b) How would you experimentally prove that the manganese dioxide was not used up in this experiment?

c) A diagram is shown of the apparatus. Predict in this experiment the manganese dioxide is acting as a __________.

Investigating how changes in concentration affect the rate of the reaction by methods involving a colour change or turbidity

Sodium thiosulfate solution (Na$_2$S$_2$O$_3$) reacts with dilute hydrochloric acid according to the equation:

$$\text{Na}_2\text{S}_2\text{O}_3(\text{aq}) + 2\text{HCl}(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{S}_2\text{O}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l})$$

In an experiment 10cm$^3$ of sodium thiosulfate was placed in a conical flask on top of a piece of paper with a cross drawn on it. Hydrochloric acid was added until the stopcock was opened. A precipitate was produced that caused the solution to become cloudy. The precipitate was orange when it was first formed but quickly dissolved in solution, due to the precipitate formed. The precipitate causes turbidity, which is measured as a fall caused by large numbers of individual unit particles.

**Questions:**

1) Look at the equation and identify the product which causes the solution to become cloudy.

2) The experiment was repeated using different concentrations of sodium thiosulfate. The results are recorded below:

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Concentration of sodium thiosulfate (M)</th>
<th>Time taken for cross to disappear (s)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.05</td>
<td>13.07</td>
</tr>
<tr>
<td>2</td>
<td>0.1</td>
<td>29.01</td>
</tr>
<tr>
<td>3</td>
<td>0.2</td>
<td>54.95</td>
</tr>
<tr>
<td>4</td>
<td>0.5</td>
<td></td>
</tr>
</tbody>
</table>

3) Calculate the rate for the rate of reaction for experiment 1 using the equation: rate = __________. Give your answer to 3 significant figures.

4) Identify three variables that must be kept constant to make this a fair test.

5) From the results of the experiment state the effect of increasing the concentration of sodium thiosulfate solution on the rate of the reaction.

6) Explain in terms of collision theory why increasing the concentration of sodium thiosulfate solution has this effect.

7) From and explain one change which could be made to this experiment to get more accurate results.

Reversible reactions and dynamic equilibrium

**Reversible reactions**

Some chemical reactions are reversible. This means that once the products have been made from the reactants, the products can react to reform the reactants. The $\leftrightarrow$ arrows are used to show that the reaction is reversible.
Reversible reactions and dynamic equilibrium

For example, arsine copper sulfate (which is white) reacts with water to form hydrated copper sulfate (which is blue). However, this reaction is reversible as when the blue hydrated copper sulfate is heated it breaks back down into white arsine copper sulfate and water (Figure 6.34).

\[ \text{arsine copper sulfate} + \text{water} \rightleftharpoons \text{hydrated copper sulfate} \]

\[ \text{CuSO}_4 \cdot 5\text{H}_2\text{O} \rightleftharpoons \text{CuSO}_4 + 5\text{H}_2\text{O} \]

When arsine chloride (NHCl) is heated it breaks down into ammonia (NH₃) and hydrogen chloride (HCl). However, when cooled the ammonia can react with the hydrogen chloride to reform arsine chloride (Figure 6.35).

\[ \text{arsine chloride} \rightleftharpoons \text{ammonia} + \text{hydrogen chloride} \]

\[ \text{NHCl} \rightleftharpoons \text{NH}_3 + \text{HCl} \]

Dynamic equilibrium

If a reversible reaction takes place in a closed system, that is, it is a closed system where no substances can get in or out, a dynamic equilibrium is reached. This happens when both the forward and reverse reactions are taking place simultaneously and at exactly the same rate of reaction.

A good analogy of a system in dynamic equilibrium is an athlete pushing up an elevator that is moving down (Figure 6.36). The athlete on the elevator is in a state of dynamic equilibrium if he or she moves at the same speed as the elevator moves down.

KEY TERMS

- Closed system: A system where no substances can get in or out.
- Dynamic equilibrium: A system where both the forward and reverse reactions are taking place simultaneously and at the same rate.

Show you can...

- Write a key term you know for the reaction. For example, NH₃ + HCl → NH₄Cl + HCl

<table>
<thead>
<tr>
<th>Table 6.4</th>
</tr>
</thead>
<tbody>
<tr>
<td>Changing the position of an equilibrium</td>
</tr>
<tr>
<td>Changing the position of an equilibrium</td>
</tr>
</tbody>
</table>

For example, if the conditions of a given chemical reaction are changed, the position of the equilibrium changes, and so the relative amount of reactants and products changes. Similar to the athlete on the elevator, if the speed of the elevator moving down was increased, then the system would no longer be in dynamic equilibrium.
equilibrium and the athlete would start moving down the escalator. If the athlete responded to this change by walking or running faster, a new equilibrium would be reached when she matches the speed of the escalator (Figure 6.37). By the time she reaches this new equilibrium, her position will have moved and she would be near the bottom of the escalator and her position will have moved to the left.

Chemical reactions at dynamic equilibrium respond to changes as well. L. E. Adelstein's principle states that:

If a change is made to the conditions of a system at equilibrium, the position of the equilibrium moves to oppose that change in conditions.

The effect of changing concentration
If the concentration of a chemical in an equilibrium is changed, the position of the equilibrium will move to oppose that change. If more of a chemical is added, the equilibrium position moves to remove it. If some of a chemical is removed, the equilibrium position moves to make more of it.

Table 6.6 illustrates how changes in concentration affect the position of an equilibrium of the form:

\[ \text{reactants} \rightleftharpoons \text{products} \]

<table>
<thead>
<tr>
<th>Initial concentration</th>
<th>Increased concentration</th>
<th>Reaction shift</th>
<th>Decreased concentration</th>
<th>Reaction shift</th>
</tr>
</thead>
<tbody>
<tr>
<td>Reactant A</td>
<td></td>
<td>Move right</td>
<td>Move left</td>
<td></td>
</tr>
<tr>
<td>Product B</td>
<td></td>
<td>Move left</td>
<td>Move right</td>
<td></td>
</tr>
</tbody>
</table>

However, reactions where one product is formed at lower temperatures are not usually carried out at low temperature. This is because they would be far too slow. A compromise temperature is usually used that gives a reasonable yield of product but is not too slow.

Example
Ammonia is made by reacting nitrogen with hydrogen. What would happen to the amount of ammonia formed if the concentration of nitrogen was increased?

Answer
The amount of ammonia formed would increase because the equilibrium position moves right if the concentration of nitrogen is increased.

Example
Hydrogen can be made by reacting methane with steam. What would happen to the amount of hydrogen formed if the carbon monoxide was removed?

Answer
The amount of hydrogen formed would decrease because the equilibrium position moves left if the carbon monoxide is removed.
### Table 6.1

<table>
<thead>
<tr>
<th>Table 6.1: Reversible reactions and dynamic equilibria</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Equilibrium Reactions</strong></td>
</tr>
<tr>
<td><strong>Forward Reaction</strong></td>
</tr>
<tr>
<td><strong>Products</strong></td>
</tr>
<tr>
<td><strong>Equilibrium Constant (K)</strong></td>
</tr>
<tr>
<td><strong>Law of Mass Action</strong></td>
</tr>
<tr>
<td><strong>Chemical Equation</strong></td>
</tr>
<tr>
<td><strong>Temperature</strong></td>
</tr>
<tr>
<td><strong>Pressure</strong></td>
</tr>
<tr>
<td><strong>Concentration</strong></td>
</tr>
</tbody>
</table>

### The effect of catalysts

Catalysts have no effect on the position of an equilibrium. However, the catalyst does increase the rate of forward and backward reactions, both by the same amount. This means that the system reaches equilibrium faster and the product is formed faster.

### Example

Ammonia is made by reacting nitrogen with hydrogen. What would happen to the amount of ammonia formed if the pressure was increased?

**Answer**

- The equilibrium position moves to the right with fewer gas molecules to decrease the pressure.
- More ammonia is produced because the equilibrium position moves right.

### Example

Hydrogen can be made by reaction of methanol with steam. What would happen to the amount of hydrogen formed if the pressure was increased?

**Answer**

- The equilibrium position moves to the side with fewer gas molecules to decrease the pressure.
- Less hydrogen is produced because the equilibrium position moves left.

### Example

Hydrogen gas reacts with chlorine gas to form hydrogen chloride gas. What would happen to the amount of hydrogen chloride formed if the pressure was increased?

**Answer**

- The equilibrium position does not move because there are the same number of gas molecules on both sides of the equation.
- The amount of hydrogen chloride formed remains the same.
Chapter review questions

1. Carbon monoxide reacts with steam to form carbon dioxide and hydrogen in a reversible reaction. The reaction is carried out in a catalytic reactor in order to make hydrogen. Write the balanced chemical equation. The pressure in the reactor is maintained at 1 atm. Carbon monoxide is fed into the reactor at a rate of 0.1 mol/min. Assuming that the reaction goes to completion, calculate the rate of production of hydrogen at steady state.

2. A substance undergoes a series of reactions as shown below:

\[ \text{A} \rightarrow \text{B} \rightarrow \text{C} \rightarrow \text{D} \]

If the rate of reaction for each step is represented by \( r_1, r_2, r_3, \) and \( r_4 \) respectively, what is the overall rate of reaction? Explain your answer.

3. A chemical reaction proceeds as follows:

\[ \text{A} + \text{B} \rightarrow \text{C} \]

If the concentration of \( \text{A} \) is doubled, and the concentration of \( \text{B} \) is halved, how does the rate of the reaction change? Explain your answer.

4. A reaction is carried out at different temperatures. The rate of the reaction is found to be first-order with respect to \( \text{A} \) and second-order with respect to \( \text{B} \). Write the rate law for the reaction.

5. A reaction is carried out in a closed system at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

6. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

7. A reaction is carried out in a closed system at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

8. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

9. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

10. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

11. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

12. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

13. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

14. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

15. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

16. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

17. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

18. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

19. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

20. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

21. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

22. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

23. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

24. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

25. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

26. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

27. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

28. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

29. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

30. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

31. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

32. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

33. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

34. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

35. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

36. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

37. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

38. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

39. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

40. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

41. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.

42. A reaction is carried out at a pressure of 1 atm. If the pressure is increased to 2 atm, how does the rate of the reaction change? Explain your answer.

43. A reaction is carried out at a temperature of 300 K. If the temperature is increased to 600 K, how does the rate of the reaction change? Explain your answer.
Practice questions

1. Which one of the following is used to increase the rate of an exothermic reaction?
   - [ ] catalyst
   - [ ] oxidizing agent
   - [ ] reducing agent
   - [ ] increased temperature
   - [ ] reduced temperature

2. In which of the following experiments will the rate of reaction be quickest at the start of the reaction?
   - [ ] zinc powder reacting with an excess of 3 M HCl at 30°C
   - [ ] zinc powder reacting with an excess of 1 M HCl at 30°C
   - [ ] zinc powder reacting with an excess of 3 M HCl at 80°C
   - [ ] zinc powder reacting with an excess of 1 M HCl at 80°C

3. In the laboratory preparation of oxygen from hydrogen peroxide using manganese(IV) oxide as catalyst, the mass of manganese(IV) oxide was measured at various times. Which one of the following graphs best shows the experimental results?

4. In a laboratory experiment, 50g of magnesium ribbon was reacted with excess dilute hydrochloric acid at room temperature. The volume of gas produced was noted every 30 seconds. The results were plotted on the axes below as graph C.

   - [ ] What is the volume of gas produced after 60 seconds?
   - [ ] What is the rate of reaction at 30 seconds?
   - [ ] Does the reaction slow down or speed up with time?

5. What is the rate of reaction at time t = 30 seconds?

6. Calculate the rate of reaction at time t = 0 seconds.

7. Calculate the rate of reaction between time t = 0 seconds and t = 30 seconds.

8. Calculate the rate of reaction between time t = 30 seconds and t = 60 seconds.

Working scientifically: Presenting information and data in a scientific way

**Recording results**

During experimental activities you need often record results in a table. When drawing tables and recording data ensure that:
- the table is clearly divided into columns and rows
- there are headings for each column or row
- there are subheadings for each column and row
- there are units for each column and row
- there is room for repeat measurements and averages

**Questions**

1. 50cm³ of hydrogen peroxide and 1.5g of manganese dioxide were mixed and heated to 30°C. The volume of oxygen collected from the reaction at 15 second intervals is:
   - [ ] 10 cm³
   - [ ] 20 cm³
   - [ ] 30 cm³
   - [ ] 40 cm³

2. State and explain the effect of an increase in pressure on the yield of ethanol.

3. State and explain the effect of an increase in temperature on the yield of ethanol.

4. What is a catalyst?

5. State the effect of the catalyst on the yield of ethanol.

6. Assuming 80% yield, what mass of ethanol is needed to produce 1 L of ethanol?

7. Nitrogen oxide (NO) is a known gas which can form from the combustion of nitrogen tetraoxide (N₂O₄) in a dynamic equilibrium. The energy change for the forward reaction is exothermic.

   \[2NO₂(g) \rightleftharpoons N₂O₄(g)\]

   What is meant by the term dynamic equilibrium? (2 marks)

8. Explain what is observed when the pressure on the equilibrium mixture is increased. (2 marks)

9. Explain what is observed when the temperature of the equilibrium mixture is reduced. (2 marks)

Plotting graphs

When carrying out experiments you need to be able to translate data from one form to another. Think of this procedure using data from a table to draw a graph. A graph is an illustration of how two variables vary one another. When drawing a graph remember that:

- the independent variable is placed on the x-axis, while the dependent variable is placed on the y-axis
- the independent variable should be divided for easy reading, meeting the most significant one of the table as the graphs area. The scale should be selected so that the results are not exaggerated or reduced.
Organic chemistry is the study of compounds containing carbon. The number of organic compounds that exist is greater than all other compounds added together. This is because carbon is able to form bonds to other carbon atoms very easily and form chains and rings. Many medicines, detergents, clothing fibres, solvents and polymers are organic molecules and many of these are made from chemicals that we find in crude oil. Our own bodies and food also contain many organic molecules.

This chapter covers specification points a, b, and c only and is called Organic Chemistry.

In chemistry, carbon compounds as fuels and textiles, reactions of alkenes and alcohols, and synthetic and naturally occurring polymers.

7 Organic chemistry

Questions

A) 5g of some calcium carbonate (carbonate salt) were placed into a vessel. The vessel was then exposed to air and the balance reading recorded every minute. The results were recorded and the graph shown below was drawn.

- What are the y-axis units from the graph?
- What is the independent variable in the experiment?
- At what time does the reaction stop?
- What is the mass of the vessel and contents at time 0 (minutes)?

Tip: You need to judge the points of change in the graph from the red dot on the graph.

A line of best fit should be drawn when judging the position of the two should be approximately the same number of data points on each side of the line. Read off the tangents to simply connect the first and last points. The best fit line could be straight or a curve. Ignore any erroneous results when drawing your line.
Crude oil and alkanes

Previously you could have learned:

- Crude oil, coal and natural gas are fossil fuels.
- Fossil fuels are substances that were in-oil releasing a lot of thermal energy.
- Non-metal atoms bond to each other by sharing electrons; one covalent bond is made up of two shared electrons.
- Carbon atoms make four covalent bonds to molecules.
- Mixtures of alkanes liquids with different boiling points can be separated by fractional distillation.
- Polymers are long chain molecules.

Test yourself on prior knowledge:

- What are fossil fuels?
- Give three examples of fossil fuels.
- What type of bonds are made when a metal atom bonds with a non-metal atom?
- How many covalent bonds do carbon atoms make?
- What is a polymer?
- How are mixtures of alkanes liquids separated?

Crude oil and alkanes

Crude oil is a fossil fuel that is found underground in rock. It was formed over millions of years from the remains of sea creatures. These creatures were mainly plants that were buried in mud at the bottom of the ocean (Figure 7.1).

KEY TERMS

- Fossil fuels: A resource that cannot be replaced once it has been used.
- Biomass: A resource made from living or recently living organisms.
- Hydrocarbons: A compound containing hydrogen and carbon only.

Crude oil is a mixture of many different compounds. Most of these compounds are hydrocarbons. Hydrocarbons are compounds that contain hydrogen and carbon only. Most of the hydrocarbons in crude oil are alkanes.

Alkanes

Alkanes are a family of saturated hydrocarbons. Saturated molecules are those that contain single covalent bonds. The structure of the first four alkanes are shown in Table 7.1. The table includes the displayed formula which shows all the atoms and all the bonds in each molecule.

<table>
<thead>
<tr>
<th>Alkane</th>
<th>Chain and ring</th>
<th>Displayed structure</th>
<th>Structural formula</th>
<th>Molecular formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methane</td>
<td>CH₄</td>
<td></td>
<td>CH₄</td>
<td></td>
</tr>
<tr>
<td>Ethane</td>
<td>C₂H₆</td>
<td></td>
<td>C₂H₆</td>
<td></td>
</tr>
<tr>
<td>Propane</td>
<td>C₃H₈</td>
<td></td>
<td>C₃H₈</td>
<td></td>
</tr>
<tr>
<td>Butane</td>
<td>C₄H₁₀</td>
<td></td>
<td>C₄H₁₀</td>
<td></td>
</tr>
</tbody>
</table>

The names of organic compounds are made up of two parts. The first part indicates which homologous series the molecule belongs to (Figure 7.2).
The alkanes are an example of a homologous series. A homologous series is a family of compounds with
- the same general formula
- the same functional group
- similar chemical properties.

Alkanes do not contain a functional group. Functional groups will be studied later in the chapter.

### Fractional distillation of crude oil

For crude oil to be useful, the hydrocarbons it contains have to be separated. The hydrocarbons have different boiling points and this difference is used to separate them by fractional distillation at an oil refinery. This process separates the hydrocarbons into fractions. A fraction is a mixture of molecules with similar boiling points. In each fraction, the hydrocarbons contain a similar number of carbon atoms.

The crude oil is heated and vaporized. The vaporized crude oil enters the fractionating tower that is hotter at the bottom and cooler at the top. The hydrocarbons cool as they rise up the tower and condense at different heights because they have different boiling points. Hydrocarbons with large molecules are collected as liquids near the bottom of the tower while those with small molecules collect at the top (Figure 7.3).

### The use of alkanes as fuels

The main use of alkanes from crude oil is as fuels. Alkanes are good fuels because they release a lot of energy when they burn.

When alkanes burn, they react with oxygen. Complete combustion takes place if there is a good supply of oxygen from the air. The carbon atoms in the alkanes are oxidized, combining with the oxygen to form carbon dioxide. The hydrogen atoms in the alkanes are also oxidized, combining with the oxygen to form water.

\[ \text{Incomplete combustion: } \text{alkane} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water} \]

If there is a poor supply of oxygen when alkanes burn, incomplete combustion can happen which forms water along with carbon monoxide and/or carbon (in the form of soot). The carbon monoxide formed is toxic and the soot causes a smoky flame.

\[ \text{Incomplete combustion: } \text{alkane} + \text{oxygen} \rightarrow \text{carbon monoxide} + \text{water} \]

When writing balanced equations for the complete combustion of alkanes:
- Balance the C atoms for each C atom in the alkanes there will be one \( \text{CO}_2 \) molecule formed.
- Balance the H atoms for every two H atoms in the alkanes there will be one \( \text{H}_2 \) molecule formed.
- Count the number of O atoms in the \( \text{CO}_2 \) and \( \text{H}_2 \)O; the number of O atoms will be half this number.
- If the number of \( \text{O}_2 \) molecules has a half in it, double all the balancing numbers to get rid of the half.

### Example

Write a balanced equation for the complete combustion of methane, \( \text{CH}_4 \).

\[ \text{Complete combustion equation: } \text{methane} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water} \]

\[ \text{balanced equation: } \text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2 \text{O} \]

\[ \text{Reactions: } \text{CH}_4 + 2 \text{O}_2 \rightarrow 2 \text{H}_2 \text{O} + \text{CO}_2 \]
The properties of alkanes depend on the size of the molecules and this affects their use as fuels (Figure 7.4). The flammability of a fuel is how easily it catches fire. A flammable fuel catches fire easily. The viscosity of a liquid is how easily it flows. A runny liquid has a low viscosity while a thick, slow moving liquid (e.g. syrup) has a high viscosity.

**Example**

Write a balanced equation for the complete combustion of butane, C₄H₁₀.

**Answer**

Current equation: butane + oxygen → carbon dioxide + water

Balanced equation: 2C₄H₁₀ + 13O₂ → 8CO₂ + 10H₂O

**Test yourself**

1. Describe how grade oil was formed.
2. Explain why crude oil can be described as an ancient source of biomass.
3. Create a list of hydrocarbons. Explain what this means.
4. Create an oil mixture of hydrocarbons. They are separated by fractional distillation at an oil refinery.
5. Which is a hydrocarbon?
6. What property of the hydrocarbons allows them to be separated in this way?
7. State how fractional distillation separates the hydrocarbons.
8. Name two alkanes containing six carbon atoms.
9. Alkanes are saturated hydrocarbons. What does saturated mean in this context?
10. Give the molecular formula of methane.
11. Show the skeletal formula of benzene.
12. Pentane and decane are both alkanes. Pentane has the formula C₅H₁₂. Decane has the formula C₁₀H₂₂. Which one is the highest boiling point? Which one is the most flammable? Which one is the most viscous?
13. Write a balanced equation for the complete combustion of the following alkanes.
   - pentane, C₅H₁₂
   - ethane, C₂H₆

**Cracking**

Shorter alkanes are in very high demand as fuels but longer alkanes are in less demand. This means that there is a surplus of the longer alkanes from the fractional distillation of crude oil. These longer alkanes can be broken down into the shorter alkanes that are in higher demand by a process called cracking. This process also produces unsaturated hydrocarbons called alkenes which can be used as a starting material to make many other substances such as polymers and medicines (Figure 7.5).

**Figure 7.3** An example of a cracking reaction.
Cracking can be done in a number of ways. Two methods of cracking are:
- Catalytic cracking: treat the alkenes to reform them and then pass them over a heat catalyst.
- Steam cracking: treat the alkenes to vaporize them, mix them with steam and then heat them to very high temperature.

Cracking is a thermal decomposition reaction because the alkenes are broken down into smaller molecules using heat.

## Alkenes

**Alkenes** are a homologous series of unsaturated hydrocarbons with the general formula \( C_nH_{2n} \).

**Unsaturated** refers to the compound's organic chemistry, which contains one or more double covalent bonds.

### Table 2.2 The first four alkenes

<table>
<thead>
<tr>
<th>Alkene</th>
<th>Structural formula</th>
<th>Chemical formula</th>
<th>Molecular mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ethene</td>
<td>( \ce{H2=CH2} )</td>
<td>( \ce{C2H4} )</td>
<td>28 g/mol</td>
</tr>
<tr>
<td>Propene</td>
<td>( \ce{H3C=CH2} )</td>
<td>( \ce{C3H6} )</td>
<td>42 g/mol</td>
</tr>
<tr>
<td>Butene</td>
<td>( \ce{H3C=CH(CH3)} )</td>
<td>( \ce{C4H8} )</td>
<td>56 g/mol</td>
</tr>
<tr>
<td>Pentene</td>
<td>( \ce{H3C=CH(CH2CH3)} )</td>
<td>( \ce{C5H10} )</td>
<td>70 g/mol</td>
</tr>
</tbody>
</table>

When drawing organic molecules, it does not usually matter what angle you draw the atoms at as long as you draw the molecules.

When drawing organic molecules, they are three-dimensional and we are representing them in two dimensions on paper and so we diagram will represent them as they actually are. For example, each of the four structures in Figure 2.7 is a correct displayed formula of propene (\( \ce{C3H6} \)).

## Alkenes contain two fewer hydrogen atoms than alkanes with the same number of carbon atoms. This is due to the double bond between two carbon atoms. The general formula of alkenes therefore is \( C_nH_{2n} \). For example, propene has 3 carbon atoms (\( n = 3 \)) and so must contain 6 hydrogen atoms (\( 2 \times 3 = 6 \)) and so has the formula \( \ce{C3H6} \).

The functional group in alkenes is the \( \ce{C=C} \) double bond. A functional group is an atom or group of atoms that is responsible for most of the chemical reactions of a compound.

### Reactions of alkenes

Alkenes are more reactive than alkanes because they contain a \( \ce{C=C} \) double bond. Most of the reactions of alkenes are **addition reactions**. In these reactions the \( \ce{C=C} \) double bond breaks open and atoms add onto the two carbon atoms. The \( \ce{C=C} \) double bond becomes a \( \ce{C=O} \) single bond as a saturated molecule is produced (Figure 7.7).

![Figure 2.7 Reaction of alkenes with halogen](image)

**Alkene molecule** molecule that reacts with alkenes. The \( \ce{C=C} \) double bond breaks open and the alkene adds onto the \( \ce{C=C} \) double bond.

**Reactions of alkenes with halogens**

The halogens chlorine (\( \ce{Cl2} \)), bromine (\( \ce{Br2} \)) and iodine (\( \ce{I2} \)) all react readily with alkenes. In each case, the \( \ce{C=C} \) double bond opens up and one halogen atom adds onto each of the carbon atoms in the double bond. Some examples are shown in Figure 7.8.

![Figure 7.8 Reaction of alkenes with halogen](image)

The reaction with bromine is used to test for the presence of \( \ce{C=C} \) double bonds in compounds. Bromine water is a solution of bromine in...
Reaction of alkenes with hydrogen
Alkenes react with hydrogen at 100°C in the presence of a nickel catalyst to form alkanes. In this reaction, hydrogen atoms add onto the carbon atoms in the C=C double bond. This converts an unsaturated alkene into a saturated alkane (Figure 7.9).

The reaction is called **hydrogenation** because hydrogen is being added across the C=C double bond. The metal, nickel, is often used as the catalyst.

**Key terms**
- Hydrogenation: Addition of hydrogen
- Hydration: Addition of water

**Test yourself**
- Alkenes can be converted to alkanes by **hydrogenation**.
- The reaction between alkenes and **hydrogen** is catalysed by nickel.
- The product of the reaction between alkenes and **hydrogen** is an alkane.

**Share this**
- The diagram shows some of the reactions of alkenes.
- The reaction shown is the addition of hydrogen across the C=C double bond.
- The product of the reaction is an alkane, which is a saturated hydrocarbon.

---

**Reaction of alkenes with oxygen**
Alkenes will also burn in air like alkanes do. However, we do not usually burn alkenes for the same reasons.

1. **Alkenes tend to burn in air with bright flames due to incomplete combustion.**
2. **Alkenes are very valuable because they can be used to make polymers or as a starting material for many other chemicals.**
Alcohols, carboxylic acids and esters

Alcohols
Alcohols are a homologous series of compounds containing the functional group –OH. The first four alcohols are shown in Table 7.5. The names of alcohols end in –ol.

<table>
<thead>
<tr>
<th>Alcohol</th>
<th>Molecular formula</th>
<th>Structural formula</th>
<th>Molecular formula</th>
<th>Structural formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ethanol</td>
<td>CH₃CH₂OH</td>
<td>CH₃CH₂OH</td>
<td>C₂H₆O</td>
<td>CH₃CH₂OH</td>
</tr>
<tr>
<td>Propanol</td>
<td>CH₃CH₂CH₂OH</td>
<td>CH₃CH₂CH₂OH</td>
<td>C₃H₇O₂</td>
<td>CH₃CH₂CH₂OH</td>
</tr>
<tr>
<td>Butanol</td>
<td>C₃H₇OH</td>
<td>C₃H₇OH</td>
<td>C₄H₉O₂</td>
<td>C₃H₇OH</td>
</tr>
</tbody>
</table>

Making ethanol
Ethanol is made by the fermentation of sugars in these conditions:
- the sugars are dissolved in water
- yeast is added
- the mixture is kept in a warm place at about 30°C
- air is kept out of the mixture.

Reactions of alcohols
Combustion reactions
Alcohols burn in oxygen. In a good supply of oxygen, complete combustion takes place to form carbon dioxide and water, for example:

- ethanol + oxygen → carbon dioxide + water
- C₂H₅OH + 3O₂ → 2CO₂ + 3H₂O
- propanol + oxygen → carbon dioxide + water
- C₃H₇OH + 5O₂ → 3CO₂ + 3H₂O
- butanol + oxygen → carbon dioxide + water
- C₄H₉OH + 7O₂ → 4CO₂ + 7H₂O

Mild oxidation of alcohols
When alcohols are burned, they form carbon dioxide and water if combustion is complete. If dilute solutions of alcohols in water are left standing in air, the alcohols will only partially oxidise and form carboxylic acids (Figure 7.12).

- Ethanol + oxygen → ethanoic acid + water
- C₂H₅OH + O₂ → CH₃CO₂H + H₂O

- Ethanoic acid + sodium hydroxide → ethanol + water + sodium
- CH₃CO₂H + NaOH → CH₃CO₂Na + H₂O

- Ethanol + oxygen + sodium hydroxide → ethanoic acid + water + sodium
- C₂H₅OH + O₂ + NaOH → CH₃CO₂Na + H₂O

- Ethanoic acid + ethanoic acid + water + sodium hydroxide → ethanol + sodium + water
- CH₃CO₂H + CH₃CO₂H + NaOH → C₂H₅OH + Na₂CO₂ + H₂O

- Ethanoic acid + ethanoic acid + water + sodium hydroxide → ethanoic acid + sodium + water
- CH₃CO₂H + CH₃CO₂H + NaOH → 2CH₃CO₂Na + H₂O

- Ethanoic acid + ethanoic acid + water + sodium hydroxide → ethanoic acid + sodium + water
- CH₃CO₂H + CH₃CO₂H + NaOH → 2CH₃CO₂Na + H₂O
Uses of alcohols

Alcohols are very useful substances. Some of their uses are shown below.

- As fuels: Alcohols burn very well and can be used as fuels. Compare gas stoves, which use alcohols as fuels. In every litre of petrol sold in the UK, 50 cm³ (15%) of the fuel is actually ethanol. In Brazil, many cars run on ethanol as a fuel instead of petrol.
- As solvents: Alcohols are very good solvents. For example, many solutions of medicines and perfumes are made with alcohols because these substances do not dissolve in water.
- In alcoholic drinks: Ethanol is the main alcohol in alcoholic drinks. Alcoholic drinks are made by fermentation of crops such as grapes (to make wine), apples (to make cider) or malted barley (to make beer). In each case, ethanol is produced by yeast in fermentation.

**Key Terms**

Carboxylic acids are a homologous series of compounds containing the functional group –COOH. The first four carboxylic acids are shown in Table 7.6. The names of carboxylic acids end in -oic acid.

<table>
<thead>
<tr>
<th>Carboxylic acid</th>
<th>Structural formula</th>
<th>Molecular formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methanoic acid</td>
<td>HCOOH</td>
<td>CH₃COOH</td>
</tr>
<tr>
<td>Ethanoic acid</td>
<td>CH₃COOH</td>
<td>CH₃COOH</td>
</tr>
<tr>
<td>Propanoic acid</td>
<td>CH₃CH₂COOH</td>
<td>CH₃CH₂COOH</td>
</tr>
<tr>
<td>Butanoic acid</td>
<td>CH₃CH₂CH₂COOH</td>
<td>CH₃CH₂CH₂COOH</td>
</tr>
</tbody>
</table>

Many everyday substances also contain the carboxylic acid functional group (Figure 7.16).

Properties of carboxylic acids

*Weak acids*

Carboxylic acids dissolve in water to produce acidic solutions (Figure 7.17). Acids are substances that produce H⁺ ions when added to water.

\[
\text{RCOOH} \rightarrow \text{RCOO}^- + \text{H}^+
\]

Carboxylic acids are weak acids, which means only a small fraction of the molecules break down into ions when added to water (see page 109). As they are only weak acids, carboxylic acids do not cause as much harm as when we eat or drive them in small amounts.

**Reaction with carbonates**

Acids react with metal carbonates to form a salt, carbon dioxide and water. This means that there is fizzing when metal carbonates are added to aqueous solutions of carboxylic acids. However, the reaction is slow as carboxylic acids are weak acids.

**Reaction with alcohols**

Carboxylic acids react with alcohols in the presence of an acid catalyst to form compounds called esters. For example, ethanoic acid reacts with ethanol to form the ester ethyl ethanoate which is used as the solvent in nail varnish (Figures 7.18 and 7.19).
Substances that dissolve in water to form a solution that turns a specific indicator solution one color are called **substances**. Some examples of these are **alcohols**, **esters**, and **carboxylic acids**.

**Alcohols** contain the functional group #-OH#. **Esters** contain the functional group #-COO#. **Carboxylic acids** contain the functional group #-COOH#. These functional groups are important in various chemical reactions and play a crucial role in the properties of the substances they are part of.

# Test yourself #

1. **Look at the following molecules A to H:**

   ![Molecules A to H](image)

   a. Which of these molecules are alcohols? 
   b. Which of these molecules are carboxylic acids? 
   c. Which of these molecules are esters?

2. **Look at each pair of substances:****

   a. Water and an alcohol.
   b. Water and an ester.
   c. Water and a carboxylic acid.

   Which pair is most likely to react and form a product?

3. **Look at each pair of substances:****

   a. Water and an alcohol.
   b. Water and an ester.
   c. Water and a carboxylic acid.

   Which pair is least likely to react and form a product?

4. **Look at each pair of substances:**

   a. Water and an alcohol.
   b. Water and an ester.
   c. Water and a carboxylic acid.

   Which pair is likely to react and form a product if a catalyst is added?

5. **Look at the following reaction:****

   ![Reaction](image)

   a. Identify the reactants.
   b. Identify the products.
   c. Balance the equation.

6. **Look at the following reaction:****

   ![Reaction](image)

   a. Identify the reactants.
   b. Identify the products.
   c. Balance the equation.

7. **Look at the following reaction:****

   ![Reaction](image)

   a. Identify the reactants.
   b. Identify the products.
   c. Balance the equation.

8. **Look at the following reaction:****

   ![Reaction](image)

   a. Identify the reactants.
   b. Identify the products.
   c. Balance the equation.

9. **Look at the following reaction:****

   ![Reaction](image)

   a. Identify the reactants.
   b. Identify the products.
   c. Balance the equation.

10. **Look at the following reaction:****

   ![Reaction](image)

   a. Identify the reactants.
   b. Identify the products.
   c. Balance the equation.

11. **Look at the following reaction:****

   ![Reaction](image)

   a. Identify the reactants.
   b. Identify the products.
   c. Balance the equation.

12. **Look at the following reaction:****

   ![Reaction](image)

   a. Identify the reactants.
   b. Identify the products.
   c. Balance the equation.

13. **Look at the following reaction:****

   ![Reaction](image)

   a. Identify the reactants.
   b. Identify the products.
   c. Balance the equation.

14. **Look at the following reaction:****

   ![Reaction](image)

   a. Identify the reactants.
   b. Identify the products.
   c. Balance the equation.

15. **Look at the following reaction:****

   ![Reaction](image)

   a. Identify the reactants.
   b. Identify the products.
   c. Balance the equation.

16. **Look at the following reaction:****

   ![Reaction](image)

   a. Identify the reactants.
   b. Identify the products.
   c. Balance the equation.

# Show you can... #

1. Write the balanced equation for the reaction between #\text{alcohol A}# and #\text{acid C}#.
2. Write the balanced equation for the reaction between #\text{alcohol A}# and #\text{ester D}#.
3. Write the balanced equation for the reaction between #\text{ester D}# and #\text{acid C}#.
Addition polymers

Paper clips can be joined together to make a long chain (Figure 7.21).

In a similar way, molecules containing a C=C double bond can react with each other in addition reactions. They join onto each other to create a long chain molecule called a polymer.

The C=C double bonds open up and the molecules join onto each other to make a long chain molecule. The exact number of molecules that join together varies, but is likely to be (several) hundred. For example, lots of ethene molecules join together to make the polymer polyethylene, better known as polythene (Figure 7.22).

The chemical equation for this reaction is shown in Figure 7.23.

The polymer is made up of a unit that repeats many times – this is known as the repeating unit. We can draw a single repeating unit or show the full structure of the polymer (Figure 7.24).

The small molecules that we join together to make a polymer are called monomers. In addition polymerisation, the monomers all contain a C=C double bond. Some more examples of monomers and the polymers that were formed are shown in Table 7.7. The name of a polymer is the word poly followed by the name of the monomer in brackets.

Table 7.7 Ethene polymers and their uses.

<table>
<thead>
<tr>
<th>Ethene polymers</th>
<th>Structure of polymer</th>
<th>Uses of polymer</th>
</tr>
</thead>
<tbody>
<tr>
<td>Polyethylene</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Polypropene</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Polystyrene</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

When monomers with C=C double bonds join together in addition polymerisation, no other product is formed.

TIP

Typically several hundred or even thousands of monomers join together in each chain.

KEY TERMS

Monomer (Gr. molecules that are joined together to make a polymer.

Polymer (Gr. poly, many, + meros, part) made from joining lots of monomers together.

Addition polymerisation (Gr. addition, to join; polymer, chain) where many small molecules are joined together to make a long chain molecule and nothing else is produced.

repeating unit of polymer |
structure of polymer |

(a) What is a monomer?
(b) What is a polymer?
(c) What functional group do monomers that form addition polymers contain?
(d) What is a monomer?
(e) What is a polymer?
(f) What is addition polymerisation?
(g) What is the structure of ethene shown?
(h) Name the repeating unit of polymer.
(i) How many monomers are in the polymer shown?
Condensation polymers

Condensation polymerisation (also known as condensation) is a reaction where many small molecules are joined together to make a polymer. As the monomers join together, small molecules such as water are released which is why this is called condensation polymerisation.

There are many examples of condensation polymers. Some are synthetic ones including polymers such as terylene, and polyamides such as nylon and Kevlar. There are also many naturally occurring condensation polymers such as starch, cellulose and proteins and these are studied on page 193.

Polymers

Polymers are made when molecules with two carboxylic acid functional groups react with molecules with two alcohol functional groups. The carboxylic acid groups react with the alcohol groups to form ester linkages (–CO–) and give off water (Figure 7.25).

The overall equation for this reaction can be written as in Figure 7.26.

The structure of the repeating unit and of the polymer can be represented as in Figure 7.27.

Polymers can also be formed from monomers that each contain one carboxylic acid group and one alcohol group (Figure 7.28).

Key

A group of atoms between the two functional groups

Figure 7.28 Making polymers from one monomer.

Polyamides

Polyamides are made when molecules with two carboxylic acid functional groups react with molecules with two amine functional groups. The carboxylic acid groups react with the amine groups to form amide linkages (–CONH–) and give off water (Figure 7.29).

The overall equation for this reaction can be written as in Figure 7.30.

Polyamides can also be formed from monomers that each contain one carboxylic acid group and one amine group (Figure 7.31).

Table 7.8 gives some specific examples of condensation polymers formed in this way.
### Table 3.8: Some condensation polymers and their monomers.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Polymer</th>
<th>Monomer</th>
<th>Monomer</th>
<th>Monomer</th>
<th>Monomer</th>
</tr>
</thead>
<tbody>
<tr>
<td>H₂O + CH₂=CH–CO–OH</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>H₂O + CH₂=CH–COO⁻</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>H₂O + CH₂=CH–CO–OH</td>
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<td></td>
</tr>
<tr>
<td>H₂O + CH₂=CH–COO⁻</td>
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</tr>
<tr>
<td>H₂O + CH₂=CH–CO–OH</td>
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</tr>
<tr>
<td>H₂O + CH₂=CH–COO⁻</td>
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</tr>
<tr>
<td>H₂O + CH₂=CH–CO–OH</td>
<td></td>
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<td></td>
<td></td>
</tr>
<tr>
<td>H₂O + CH₂=CH–COO⁻</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Naturally occurring polymers

Proteins

Proteins are polymer molecules made from lots of different amino acids joined together. There are 20 amino acids that make proteins. In each protein, these different amino acids are joined together in a different but very specific sequence.

Starch and cellulose

Starch and cellulose are both polymer molecules made from the sugar glucose. They differ in the way the glucose molecules join together. Starch, cellulose, and glycogen are all carbohydrates. Carbohydrates are biological molecules containing carbon, hydrogen, and oxygen.

DNA

The nucleus of a cell contains chromosomes which contain genetic information. This information is needed for the development and function of all living organisms and viruses. These chromosomes are made of DNA (Deoxyribonucleic acid) (Figure 7.33).

- DNA consists of two long polymer chains that are held together in the form of a double helix. Each polymer chain in DNA is made up from four different monomers called nucleotides. These nucleotides can bond together in very many different sequences. In humans more than 99% of this sequence is the same but there is some variation from one person to another.

- Groups on the side of the nucleotides hold the two strands of the helix together (Figure 7.36). There are four different nucleotides. They can each attract one of the other nucleotides and so work in pairs to hold the two polymer strands together in the helix. These groups are sometimes abbreviated to A, T, C, and G.

Table 7.9 shows the structure of this and some other polyguanines formed from amino acids.

<table>
<thead>
<tr>
<th>Amino acid</th>
<th>Molecular structure</th>
<th>Polymerisation unit</th>
<th>Functional groups</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alanine</td>
<td>H2N(CH3)CH2COOH</td>
<td>Alanine</td>
<td>-NH2</td>
</tr>
<tr>
<td>Glutamic</td>
<td>H2N(CH3)CH2COO(NH2+)</td>
<td>Glutamic acid</td>
<td>-COO(NH2+)</td>
</tr>
<tr>
<td>Lysine</td>
<td>H2N(CH2)4CH2COOH</td>
<td>Lysine</td>
<td>-COO+</td>
</tr>
<tr>
<td>Tryptophan</td>
<td>H2N(CH2)2CH2CH2N(CH3)COOH</td>
<td>Tryptophan</td>
<td>-COO+</td>
</tr>
</tbody>
</table>

The nucleotides in DNA are adenine (A), thymine (T), cytosine (C), and guanine (G).

Figure 7.36: DNA consists of two polymer strands held together in a double helix.
Chapter review questions

1. The structure of five molecules are shown.
   a) Which of these molecules is:
      i) an alkane
      ii) an alkyne
      iii) a carboxylic acid
      iv) a polymer?
   b) Name each of these molecules.
   c) Hexane is a saturated hydrocarbon. Hexane is an unsaturated hydrocarbon.
      a) Define the terms saturated and unsaturated.
      b) These two compounds can be distinguished using bromine water. What would happen to orange bromine water if it was added to:
         i) hexane?
         ii) hexene?
   d) Identify the functional group in each of the following molecules.
      a) There is a group which when sodium carbonate is added to molecule A:
         i) forms a solution with 
         ii) forms a solution with
         iii) forms a solution with
         iv) forms a solution with
      b) Molecule B forms a solution with water, and forms a solution with 
      c) Molecule C forms a solution with bromine water, forms a solution with 
      d) Molecule D forms a solution with a solution with 
      e) Molecule E forms a solution with 
   e) Draw the molecular formula of each of the following molecules.
      a) methanol
      b) ethanol
      c) acetic acid
      d) propene
   f) Write a balanced equation for the complete combustion of each of the following compounds.
      a) pentane
      b) pentene
      c) diethyl ether
   g) Crude oil is split into fractions by fractional distillation, explain how this process separates crude oil into fractions.
   h) Crude oil is one of the fractions produced from crude oil, it is made of long chain molecules.

3. Give two reasons why long alkanes are in less demand as fuels as short alkanes.
   a) name the process used to produce short alkanes from long alkanes.
   b) Describe one way in which this process is done.
   c) What other type of compound is produced in this process besides short alkanes?
   d) Balance the following equation for this reaction that takes place in this process.
      a) What group does this molecule process that allows it to form an addition polymer?
      b) What group, if anything, is formed when this molecule reacts to form an addition polymer?
      c) Draw the repeating unit of the polymer formed.
      d) Write a balanced equation for the formation of this polymer.

5. Terylene is a condensation polymer. It is the main polyamide used in clothes. It can be made from these two monomers.
   a) draw the repeating unit of the polymer formed.
   b) Write a balanced equation for the formation of this polymer.
   c) Why is this called a condensation polymer?
   d) Molecule A is an organic acid. A polyamide can be formed when molecule A reacts together in a condensation polymerization reaction.
      a) Draw out the molecule and draw a circle around the acid functional group.
      b) Draw a square around the carboxylic acid functional group.
      c) Draw the repeating unit of the polyamide formed.
      d) Write a balanced equation for the formation of this polyamide.
   e) The molecule shown here is an alkene with the molecular formula .
      a) The molecular formula for the reaction of this alkene with each of the substances shown is illustrated below for each of the organic molecules.
         i) bromine
         ii) ozone
         iii) hydrogen
Practice questions

1. What type of reaction takes place when CH₂OH is converted into C₆H₁₂O₆? (1 mark)
   a. condensation
   b. oxidation
   c. reduction
   d. polymerization

2. Write the molecular formula of ethylene. (1 mark)
   a. C₂H₄
   b. C₂H₂
   c. C₃H₆
   d. C₃H₈

3. Which of the following is an alcohol? (1 mark)
   a. butane
   b. ethane
   c. ethanol
   d. methane

4. Which one of the following molecular formula represents a compound which is a member of the same homologous series as C₆H₁₄? (1 mark)
   a. C₅H₁₁
   b. C₄H₉
   c. C₆H₁₂
   d. C₄H₁₀

5. What is the name of the compound represented by the structural formula below? (1 mark)
   a. pentane
   b. hexane
   c. heptane
   d. octane

6. What are the names of the two functional groups present in propene? (2 marks)

7. Write the balanced chemical equation for the reaction of ethylene and water to form ethanol. (2 marks)

8. Define what is meant by an unsaturated hydrocarbon. (1 mark)

9. Explain why propanol is a hydrocarbon. (1 mark)

10. What is the molecular formula of propanol? (1 mark)

11. Identify the two functional groups in propanol. (2 marks)

12. Write a balanced chemical equation for the reaction of carbon dioxide and water to form carbonic acid. (2 marks)

13. Write an equation for the formation of ethanoic acid from ethene. (2 marks)

14. What is a polymerisation reaction? (1 mark)

15. What are the two main functional groups found in proteins? (1 mark)

16. What is the general formula for a hydrocarbon? (1 mark)

17. What is the empirical formula of glucose? (1 mark)

18. What is the molecular formula of glucose? (1 mark)

19. What is the general formula for a hydrocarbon? (1 mark)

20. What is the empirical formula of glucose? (1 mark)
Working scientifically: Assessing risk in science experiments

When experiments and investigations are carried out in the laboratory, you need to decide if the experiment is safe by carrying out a risk assessment. A risk assessment is a judgement of how likely it is that someone might come to harm if a planned action is carried out and how these risks could be reduced.

A good risk assessment includes:
1. A list of all the hazards in the experiment.
2. A list of the risks that the hazards could cause.
3. Suitable control measures you could take which will reduce or prevent the risk.

For chemicals there should be a written hazard warning sign on the container.

**KEY TERMS**
1. **Toxic**
   - Harmful if swallowed, inhaled, or absorbed through the skin.
   - Avoid contact with skin and eyes.
2. **Corrosive**
   - Causes burns or irritation to the skin or eyes.
   - Avoid contact with skin and eyes.
3. **Explosive**
   - Can explode if ignited.
   - Avoid contact with heat, flame, or other sources of ignition.
4. **Flammable**
   - Can burn easily.
   - Avoid contact with heat, flame, or other sources of ignition.
5. **Irritant**
   - Causes irritation to the skin or eyes.
   - Avoid contact with skin and eyes.
6. **Carcinogenic**
   - Can cause cancer.
   - Avoid contact with skin and eyes.

**Hazard**
- **Chemical**
- **Physical**
- **Use**
- **Storage**
- **Transfer**
- **Disposal**

**Control Measures**
- **Personal Protection**
  - Wear appropriate protective clothing and equipment.
  - Wash hands before leaving the laboratory.
- **Safe Handling**
  - Follow the manufacturer’s instructions.
  - Keep chemicals out of reach of children.
- **Ventilation**
  - Use fume hoods when appropriate.
  - Ensure adequate ventilation.
- **Emergency Procedures**
  - Know the emergency procedures.
  - Know the location of the fire extinguisher.
- **Disposal**
  - Dispose of chemicals in the correct way.
  - Follow the manufacturer’s instructions.

**Questions**
1. The diagram shows the apparatus used to heat copper and copper sulfate in a boiling tube. Copy and complete the table to give a risk assessment for this experiment.

<table>
<thead>
<tr>
<th>Device</th>
<th>Risk</th>
<th>Control Measure</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

2. To prepare a sample of 

3. The instructions in the lab manual describe the reaction between copper and copper oxide. Write a risk assessment for this experiment.

4. What hazards and risks are associated with carrying out the experiment of a reaction in the laboratory?
Chemical analysis

Blood is a mixture of many substances, it is very important that medics can identify which substances are in blood and how much of those substances are present. Analysis is a key area of Chemistry and there are many tests and techniques that can be used to identify, measure and test the purity of a substance.

Previously you could have learned:

- A pure substance is a single element or compound.
- An element is a substance containing one type of atom only.
- A compound is a substance containing atoms of different elements joined together.
- While a pure substance is melting, freezing, boiling or condensing, the temperature remains the same.
- Coloured dyes can be analysed and separated by paper chromatography.
- There are simple tests for the gases oxygen, hydrogen and carbon dioxide.
- Ionic substances are compounds with metals combined with non-metals.

Test yourself on prior knowledge:

1. Which of the following are pure substances?
   - Air, sea water, tap water, mineral water, diamond, a strawberry, carbon dioxide
   - Here is a list of some substances: Ag, S, H2O, Ca(OH)2, CO, Co, Sn, So, KCl, K2C
   - Which of these substances are elements?
   - Which of these substances are compounds?
   - Which of the following are ionic substances?
   - Hydrogen sulfide, H2S, carbon dioxide (CO2), aluminum oxide (Al2O3), calcium bicarbonate (CaCO3), sodium nitrate (NaNO3), vitamin C (C6H8O6)

Purity, formulations and chromatography

What is a pure substance?

In everyday language, a pure substance is regarded as a natural substance that has had nothing added to it. For example, ‘pure orange juice’ is considered to be the juice taken from oranges with nothing else, such as colourings or sweeteners, added. ‘Pure water’ is considered to be water with nothing else, such as perfumes, added. Water may be regarded as a pure substance because it is pure water from the tap and nothing else is added. However, a scientist may say the word ‘pure’ in a different way and would not consider the orange juice, soap or milk to be pure substances.

A pure substance is a single element or compound. For example:
- Diamond (C) is a pure substance because it contains only carbon atoms.
- Oxygen (O2) is a pure substance because it contains only oxygen molecules.
- Glucose (C6H12O6) is a pure substance because it contains only glucose molecules.

KEY TERM

Pure substance A single element or compound that is not mixed with any other substance.
A mixture contains more than one substance. For example:
- **Orange juice** is a mixture of water molecules, citric acid and molecules, vitamin C, minerals, glucose molecules, etc. (Figure 8.3).
- **Soup** is a mixture of several simple foods from different fancy foods.
- **Mineral water** is a mixture of several substances including water, minerals, vitamins, etc.

**Melting and boiling points of pure substances and mixtures**

Pure substances melt and boil at specific temperatures. For example, water has a boiling point of 100°C, and a melting point of 0°C. While a pure substance changes state, the temperature remains constant at these values. For example, while pure water is boiling, the temperature stays at 100°C. When pure water is freezing, the temperature stays at 0°C. However, mixtures change states over a range of temperatures. Some everyday examples of this are shown in Figure 8.2.

**Test yourself**
1. Name three ways to add a mixture of mineral water is pure water.
2. Explain why some people might regard this as being pure.
3. A sample of water was found to be between -2°C and 4°C. How many different substances might it contain?
4. Three students made samples of sugar in the laboratory. The melting point of sugar is 155°C. What was the maximum possible purity of sugar if the students knew the sugar was pure?
Chromatography
What happens in paper chromatography?
Chromatography is a very useful technique that can be used to separate and analyse mixtures. There are several types of chromatography, including paper, thin layer, column and gas chromatography.

Paper chromatography is often used to analyse coloured substances. In paper chromatography:
1. A pencil line is drawn on the chromatography paper near the bottom. Pencil is used as it will not dissolve in the solvent.
2. Small amounts of the substances being analysed are placed on spots on the pencil line.
3. The paper is hung in a beaker of the solvent. The pencil line and spots must be above the level of the solvent so that the spots do not dissolve into the solvent in the beaker.
4. Over the next few minutes, the solvent soaks up the paper.
5. When the solvent is near the top, the paper is taken out of the solvent and the level that the solvent reached is marked. This is known as the solvent front.
6. The paper is left to dry.

A pure substance can only produce one spot on chromatography, whatever solvent is used. Mixtures will usually produce more than one spot, one for each substance in the mixture (figure 6.4A). It is possible to separate two substances in a mixture will move the same distance and appear as a single spot in some solvents.

How chromatography works
In each type of chromatography there is a mobile phase and a stationary phase. In paper chromatography, the stationary phase is the piece of chromatography paper and the mobile phase is a solvent.

Rf values
The ratio of the distance a substance moves to the distance moved by the solvent is called the Rf value (figure 6.5). The distance is measured to the centre of the spot.

Rf = distance moved by substance
distance moved by solvent

The Rf value for a substance is always the same in the same solvent. However, substances will have different Rf values in different solvents. Rf values can be used to identify substances.
## Identification of common gases

The gases oxygen, hydrogen, carbon dioxide, and chlorine are common gases. There is a simple chemical test to identify each one (Table 8.2).

<table>
<thead>
<tr>
<th>Gas</th>
<th>Test</th>
<th>Result</th>
</tr>
</thead>
<tbody>
<tr>
<td>Oxygen, $O_2$</td>
<td>Insert a glowing splint into a bottle of the gas. The splint might</td>
<td>Burns brightly.</td>
</tr>
<tr>
<td>Hydrogen, $H_2$</td>
<td>Insert a burning splint into a bottle of the gas. The splint might</td>
<td>Burns with a bluish tinge</td>
</tr>
<tr>
<td>Carbon dioxide, $CO_2$</td>
<td>The gas is bubbled through water or a solution of calcium hydroxide</td>
<td>Forms white deposits.</td>
</tr>
<tr>
<td>Chlorine, $Cl_2$</td>
<td>Insert a test paper into a bottle of the gas. The paper turns</td>
<td>White.</td>
</tr>
</tbody>
</table>

### Test yourself

1. Magnesium reacts with hydrochloric acid to produce a gas. This gas was collected and tested to give a simple test with a burning splint. What was the gas produced? Write a word equation for the reaction between magnesium and hydrochloric acid.
   - The gas produced is **hydrogen**.
   - A word equation for the reaction is: $Mg + 2HCl \rightarrow MgCl_2 + H_2$.

2. Copper carbonate reacts with nitric acid to produce carbon dioxide. This gas was passed through a solution of calcium hydroxide. What was the gas produced? Write a balanced equation for the reaction between copper carbonate and nitric acid.
   - The gas produced is **carbon dioxide**.
   - A balanced equation for the reaction is: $CuCO_3 + 2HNO_3 \rightarrow Cu(NO_3)_2 + CO_2 + H_2O$.

### Identification of ions by chemical and spectroscopic means

**Chemical analysis of ions**

- Compounds made from a combination of metals and non-metals are called ions. These compounds can conduct positive ions and negative ions. For example, sodium chloride contains sodium ions ($Na^+$) and chloride ions ($Cl^-$). An impure sample contains magnesium ions ($Mg^{2+}$) and sulfide ions ($S^{2-}$). Positive ions are called cations, and negative ions are called anions.

- We can test ions to identify which ions they contain. The tests for some ions are given in this chapter.
Tests for positive ions (cations)

Flame tests
Some positive ions give distinctive colours in flame tests. A simple way
to do a flame test is to dip a damp splint into the compound and then
put the splint into a varying (Barlow) flame. Table 8.3 shows the colours
produced by some common ions.

<table>
<thead>
<tr>
<th>Ion</th>
<th>Brown</th>
<th>Copper(II)</th>
<th>Purple</th>
<th>Red</th>
<th>Yellow orange</th>
<th>Light green</th>
<th>Light blue</th>
<th>Luminous green</th>
<th>Salt blue</th>
<th>Neon green</th>
<th>Clear</th>
<th>Amethyst</th>
<th>Deep violet</th>
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</thead>
<tbody>
<tr>
<td>$\text{Na}^+$</td>
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<tr>
<td>$\text{K}^+$</td>
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<tr>
<td>$\text{Al}^{3+}$</td>
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<td>$\text{Cr}^{3+}$</td>
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<td>$\text{Fe}^{3+}$</td>
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<td>$\text{Cu}^{2+}$</td>
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<td>$\text{Zn}^{2+}$</td>
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<td>$\text{Ni}^{2+}$</td>
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<td>$\text{Co}^{2+}$</td>
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<td>$\text{Mg}^{2+}$</td>
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<tr>
<td>$\text{Ca}^{2+}$</td>
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<tr>
<td>$\text{Sr}^{2+}$</td>
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<tr>
<td>$\text{Ba}^{2+}$</td>
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</tbody>
</table>

Test with sodium hydroxide solution
Sodium hydroxide solution can be used to identify some positive ions.
When sodium hydroxide solution is added to a solution of the substance
being analysed, some positive ions produce a precipitate. A precipitate
is a solid formed when two solutions are mixed.

For example, when solutions of copper(II) sulphate and sodium hydroxide
are mixed, a blue precipitate of copper(II) hydroxide is formed as the
copper ions from the copper(II) sulphate react with the hydroxide ions
from the sodium hydroxide.

- **Word equation:** Copper(II) sulphate (aq) + sodium hydroxide (aq) →
  copper(II) hydroxide (s) + sodium sulphate (aq)
- **Balanced equation:** $\text{CuSO}_4(aq) + 2\text{NaOH}(aq) → \text{Cu}(	ext{OH})_2(s) + 2\text{Na}_2\text{SO}_4(aq)$
- **Ionic equation:** $\text{Cu}^{2+}(aq) + 2\text{OH}^-(aq) → \text{Cu}(	ext{OH})_2(s)$

Solutions containing magnesium ions, calcium ions and aluminium ions all
produce a white precipitate when sodium hydroxide solution is added.
Aluminium ions can be distinguished because the white precipitate of
aluminium hydroxide dissolves when excess sodium hydroxide solution
is added, but this does not happen for calcium ions or magnesium ions.
These can be distinguished with a flame test as calcium ions give a red-
orange flame while magnesium ions do not produce any colour at all.

Tests for negative ions (anions)

Test for carbonate ions
There is a simple test for carbonate ions in a compound. When dilute
acid is added to a compound containing carbonate ions, there is Fizzing
due to the formation of carbon dioxide gas. The identity of the carbon
dioxide gas can be confirmed by testing with limewater. Carbon dioxide
forms limewater cloudy (Figure 8.1).

Most carbonate compounds are soluble in water. However, there are
some carbonate compounds including sodium carbonate and potassium
carbonate that do dissolve in water and produce solutions containing
carbonate ions.

Test for sulphate ions
There is a simple test for compounds containing sulphate ions. Dilute
hydrochloric acid followed by barium chloride solution is added to a
solution of the compound being tested. If a white precipitate forms,
the compound contains sulphate ions (Figure 8.2).

For example, when solutions of sodium sulphate and barium chloride are
mixed, a white precipitate of barium sulphate is formed as the sulphate ions
from the sodium sulphate react with the barium ions from barium chloride.

- **Word equation:** barium chloride (aq) + sodium sulphate (aq) → barium
  sulphate (s) + sodium chloride (aq)
- **Balanced equation:** $\text{BaCl}_2(aq) + \text{Na}_2\text{SO}_4(aq) → \text{BaSO}_4(s) + 2\text{NaCl}(aq)$
- **Ionic equation:** $\text{Ba}^{2+}(aq) + 2\text{Cl}^-(aq) + 2\text{Na}^+(aq) + \text{SO}_4^{2-}(aq) →$
  $\text{BaSO}_4(s) + 2\text{Na}^+(aq) + 2\text{Cl}^-(aq)$
Test for halide ions

Halide ions include chlorine ions (Cl⁻), bromide ions (Br⁻) and iodide ions (I⁻). There is a simple test for compounds containing halide ions. Sulfuric acid followed by silver nitrate solution is added to a solution of the compound being tested. Table 8.4 shows the results for different halide ions.

For example, when solutions of potassium iodide and silver nitrate are mixed, yellow precipitate of silver iodide is formed in the test tube. Then the potassium iodide meets with the silver ions from the silver nitrate.

- Net equation: silver nitrate (aq) + potassium iodide (aq) → silver iodide (s) + potassium nitrate (aq)
- Balanced equation: AgNO₃(aq) + KI(aq) → AgI(s) + KNO₃(aq)

Using the results of tests to identify an ionic compound

Example

Compound A was analysed. It gave a white frame in a flame test. When hydrochloric acid followed by barium chloride solution was added to a solution of A, a white precipitate was formed. It is the positive ion in the compound A.

1. Identify the positive ion in the compound A.
2. Identify the negative ion in the compound A.
3. Write the formula of compound A.
4. Write a balanced equation for the formation of the white precipitate.

Example

Compound B was analysed. It gave a yellow frame in a flame test. When potassium nitrate and sulfuric acid are added to a sample of B, a gas was given off that turned a flame test blue. Then the precipitate was formed. It is the positive ion in the compound B.

1. Name the gas that turns a flame test blue.
2. Name the negative ion in compound B.
3. Write the formula of compound B.
4. Write a balanced equation for the formation of the white precipitate.

Test yourself

10. Compound A was analysed. It produced a yellow frame in a flame test. When hydrochloric acid followed by silver nitrate solution was added to a solution of A, a white precipitate was produced when nitric acid followed by silver nitrate solution was added to a negative solution of E.

11. Identify the positive ion in compound A.
12. Identify the negative ion in compound A.
13. Write the formula of compound A.
14. Write a balanced equation for the formation of the white precipitate.

15. Compound B was analysed. It gave a white frame in a flame test. When hydrochloric acid followed by silver nitrate solution was added to a solution of B, a white precipitate was produced when nitric acid followed by silver nitrate solution was added to a negative solution of E.

16. Identify the positive ion in compound B.
17. Identify the negative ion in compound B.
18. Write the formula of compound B.
19. Write a balanced equation for the formation of the white precipitate.
20. Write an ionic equation for the formation of the white precipitate when carbonates are added.

TIP

A compound contains both positive and negative ions. To identify an ionic compound, both the positive and negative ions must be identified.
Investigating how paper chromatography can be used to tell the difference between different coloured solutions

When heating water for tea, it is important that the water does not contain tea in solution. In this experiment, the initial water sample was found to be coloured by the addition of tea to the water. To test the effectiveness of paper chromatography, an experiment was carried out using the water sample.

1. Place a piece of filter paper on a flat surface. Draw a straight line about 1 cm from the bottom of the paper using a pencil. Place a concentrated drop of each solution to be tested at the base of the line.
2. Place the chromatography paper into a container containing water at a depth of 1 cm. After allowing the paper to soak, remove it and shake off the excess water.
3. Lay the chromatography paper on the water bath, ensuring the ink blots are not in contact with the water. The distance travelled by each ink blot is measured and recorded.

Questions
1. Why is it necessary to draw a pencil line 1 cm from the bottom of the paper?
2. How is the water sample at a depth of 1 cm treated in the experiment?
3. What is the concentration of each solution used in the experiment and what is the effect on the chromatography?
4. The chromatogram obtained shows three bands: the top band contains the solute, the middle band is the solvent, and the bottom band is the solid. What is the significance of this?
5. What is the importance of the concentration of the solute in relation to the chromatography?
6. What is the effect of the concentration of the solute on the chromatography?
7. Why is the concentration of the solute important in relation to the chromatography?

Instrumental analysis

In addition to carrying out chemical tests in the laboratory to identify elements and compounds, there are also instrumental methods of analysis. An instrumental method of analysis involves using a range of sophisticated equipment built around a detector. Instruments can be used to identify the presence of certain components, to determine the concentration of these components, or to study their properties in more detail. In a simple experiment, a detector is used to detect the components of a sample that is passed through a column filled with a porous material that has been treated with a reagent. The components are separated into fractions based on their different properties, and the detector signals this as the fractions are collected. Instrumental methods of analysis have many uses, for example:

- to measure the concentration of a substance in a sample
- to measure the solubility of a substance
- to study the properties of a substance
- to study the properties of a sample
- to study the properties of a material

There are many advantages to using these modern, instrumental methods of analysis compared to chemical tests. These advantages include that they are:

- selective - they only use a small amount of a substance
- sensitive - they can provide a lot of valuable information about a substance
- rapid - the information can be found very quickly and accurately

Flame emission spectroscopy

When some metal ions are heated in a flame, they give a distinctive colour. For example, solutions of sodium ions give off yellow-orange light, lithium ions give off orange-red light, and calcium ions give off red-orange light. In flame tests we judge the colour by eye. Sometimes it can be difficult to judge which ion is giving off the colour. For example, some students can mix up the colour of lithium ions and calcium ions.

Flame emission spectroscopy is an example of an instrumental method of analysis. It is used for:

- to identify metal ions in solution and measure the concentration of metal ions in solution.
- to measure the concentration of metal ions in solution.

Flame emission spectroscopy is effectively a sophisticated way to do flame tests. It identifies metal ions with more certainty than flame tests and can measure the concentration of metal ions in a solution which flame tests cannot.

In flame emission spectroscopy, the light given off is measured by a spectrometer in the machine. Each metal ion gives off light at its own specific wavelength producing a line spectrum. This allows the identity of ions to be found with certainty. The emission spectra of some ions are shown in Figure 8.8, as well as the full spectrum of visible light. It can be seen, for example, that the emission spectrum of lithium ions is different from that for calcium ions even though their flames have similar colours.
The concentration of the ions in the solution can be found by measuring the intensity of a wavelength of light given out. The more intense the light it gives out, the higher the concentration of that ion. Graphs are used to show the relationship between the intensity of the light and the concentration of the ion (Figure 6.10). The intensity of the light from the sample can be measured and the graph used to find the concentration of the ions, although this is usually done directly by a computer controlling the instrument.

A solution was tested for calcium ions and found to give light with a wavelength of 435 nm, with an intensity of 3. The graph shows that this gives a concentration of 0.1 mole/1.

Use of chemical tests to identify the ions in unknown single ion compounds
A student carried out a series of tests on an ionic compound labeled A.

Questions:
1. Give an appropriate deduction for tests 2, 3, 4 and 5.
2. Name compound A.
3. Give the formula for compound A.
4. Give the formula of the white precipitate formed in test 2(a).
5. Give the formula of the blue precipitate formed in test 3(b).
Chapter review questions

1. The diagram shows some particles in air.
   a. Is air a pure substance? Explain your answer.
   b. Give the formula of all the elements shown in the diagram.
   c. Give the formula of all the compounds shown in the diagram.

2. Identify gases A, B, C and D using the results in the table.

<table>
<thead>
<tr>
<th>Test</th>
<th>Gas A</th>
<th>Gas B</th>
<th>Gas C</th>
<th>Gas D</th>
</tr>
</thead>
<tbody>
<tr>
<td>Effect on burning paper</td>
<td>Flammable</td>
<td>Inflammable</td>
<td>No effect</td>
<td>No effect</td>
</tr>
<tr>
<td>Effect on flame</td>
<td>Close to flame</td>
<td>Flame goes out</td>
<td>Flame goes out</td>
<td>Flame goes out</td>
</tr>
<tr>
<td>Effect on tomato paste</td>
<td>No effect</td>
<td>No effect</td>
<td>No effect</td>
<td>No effect</td>
</tr>
</tbody>
</table>

3. A liquid sample was analysed by paper chromatography. A spot of KNO₃ was placed on the filter paper and the paper was developed with water. The distance of migration of KNO₃ was 10 cm. After a few minutes, the paper was removed from the tank and the solvent front was at 5 cm. The results are shown below.

   a. How many substances are in the sample? [KNO₃, X, Y and Z]
   b. Which of the substances X, Y and Z are being tested?
   c. Which of the substances X, Y and Z are being tested?
   d. Which of the substances X, Y and Z are being tested?
   e. Which of the substances X, Y and Z are being tested?
   f. Why was the line drawn in pencil?
   g. Calculate the relative migration of each substance. (1 significant figure)
   h. What would happen if the 1 cm was replaced with another solvent?
   i. Turn inept is used by many people to protect them from harmful UV rays from the sun. Turn inept is a formulation, what is missing by the term formulation?

4. Some ice was taken out of a freezer and allowed to warm up. The table shows the temperature over the next 15 minutes as the ice melted.

<table>
<thead>
<tr>
<th>Time (min)</th>
<th>0</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
<th>7</th>
<th>8</th>
<th>9</th>
<th>10</th>
<th>11</th>
<th>12</th>
<th>13</th>
<th>14</th>
<th>15</th>
</tr>
</thead>
<tbody>
<tr>
<td>Temperature (°C)</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>-5</td>
<td>-10</td>
<td>-20</td>
<td>-30</td>
<td>-40</td>
<td>-50</td>
</tr>
</tbody>
</table>

   a. Plot a graph to show the temperature change over time.
   b. Write the ice point and explain your choice.
   c. Why do you think the temperature dropped so much when the ice melted?

5. The table below shows the intensity of the emission of light with wavelength 588 nm in the flame emission spectrophotometry of sodium ions at different concentrations.

<table>
<thead>
<tr>
<th>Concentration of sodium</th>
<th>Intensity</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.1 g/mL</td>
<td>0.26</td>
</tr>
<tr>
<td>0.2 g/mL</td>
<td>0.36</td>
</tr>
<tr>
<td>0.4 g/mL</td>
<td>0.52</td>
</tr>
<tr>
<td>0.6 g/mL</td>
<td>0.65</td>
</tr>
<tr>
<td>0.8 g/mL</td>
<td>0.83</td>
</tr>
<tr>
<td>1.0 g/mL</td>
<td>1.06</td>
</tr>
</tbody>
</table>

   a. Plot a graph to show how the intensity of the light varies with concentration.
   b. Draw a graph to find the concentration of sodium in a solution that emits light at 588 nm with an intensity of 0.36.
   c. Draw the advantages of using flame emission spectrophotometry methods over methods of chemical analysis for flame tests.
   d. Suppose that a sample of a unknown metal is to be analyzed. A yellow precipitate was produced when nitric acid followed by silver nitrate solution was added to a solution of KCl. Identify the negative ion in compound KCl.
   e. Write the name and formula of compound KCl.
   f. Suppose that a sample of a unknown precipitate was produced when sodium hydroxide solution was added to a solution of Na₂CO₃. This white precipitate did not dissolve when more sodium hydroxide solution was added. Faint white acid was added to the sample of Na₂CO₃. A gas was given off but turned orange slowly.

   a. What chemical test could be used to identify CO₂? Draw the possible results of the test for each compound you identify.
   b. Compound X was analyzed. A white precipitate was produced when sodium hydroxide solution was added to a solution of KNO₃. When hydrochloric acid was added to a solution of X, a white precipitate was produced. Identify the positive ion in compound X.
   c. Write the name and formula of compound X.
   d. Write a balanced equation for the formation of the brown precipitate when sodium hydroxide is added.
   e. Write a balanced equation for the formation of the white precipitate when hydrochloric acid is added.
   f. Write a balanced equation for the formation of the white precipitate when sodium hydroxide is added.
   g. Write an ionic equation for the formation of the white precipitate when sodium hydroxide is added.
Practice questions

1. A solid is thought to be pure benzoic acid. Which of the following is the best way to test to purify it? [1 mark]
   a) determine the density
   b) determine the pH
   c) determine the melting point
   d) determine the flame colour

2. Which one of the following solutions will give a coloured precipitate when a few drops of sodium hydroxide solution is added? [1 mark]
   a) sodium nitrate
   b) calcium nitrate
   c) copper nitrate
   d) sodium chloride

3. The gas produced when solid potassium carbonate reacts with hydrochloric acid causes bubbles to turn cloudy. Which of the following could be solid KOH? [1 mark]
   a) calcium carbonate
   b) calcium chloride
   c) calcium hydroxide
   d) calcium oxide

4. A test tube contains dilute nitric acid and test-tube B contains dilute hydrochloric acid. Which of the following substances can be added to each test tube and produce different observations in each? [1 mark]
   a) barium chloride
   b) copper sulphate
   c) calcium hydroxide
   d) calcium carbonate

5. The sodium chloride and silver nitrate solutions were added to a sample of the solid solution. [2 marks]
   a) Write a balanced ionic equation for the reaction of hydrochloric acid with silver nitrate solution.
   b) Explain your answer to the formation of the white precipitate in test 3.
   c) The presence of sodium ions can be determined by adding two solutions and observing the formation of a precipitate and its colour.
   d) What is an anion? [2 marks]
   e) What is meant by the term precipitation? [2 marks]
   f) Name the solution which is used to test for the presence of silver ions. [2 marks]
   g) Potassium iodide solution was mixed with zinc nitrate solution and a precipitate formed. State the colour of the precipitate. [2 marks]

6. Using the evidence from Test 1, only, name a cation which may be present in the mixture. [1 mark]
   b) Using the evidence from Test 3, only, name a cation which may be present in the mixture. [1 mark]
   c) Using the evidence from Test 1, only, name a cation which may be present in the mixture. [1 mark]
   d) Suggest the names of two compounds which may be present in the mixture. [2 marks]
   e) Write an ionic equation for the formation of the white precipitate in test 3. [2 marks]
   f) The presence of potassium ions can be determined by mixing two solutions and observing the formation of a precipitate and its colour.
   g) What is an anion? [2 marks]
   h) What is meant by the term precipitation? [2 marks]
   i) Name the solution which is used to test for the presence of silver ions. [2 marks]
   j) What is the advantage of using the method described? [2 marks]
   k) Advantages of using the method of paper chromatography. [2 marks]

7. A sports drink contains many different salts and some food colouring. In a laboratory a sports drink was analysed using paper chromatography. Describe how this experiment was carried out. [5 marks]

8. Chemical analysis

<table>
<thead>
<tr>
<th>Compound</th>
<th>Mass of Compound</th>
<th>Mass of Water</th>
<th>Mass of Dry Substances</th>
</tr>
</thead>
<tbody>
<tr>
<td>Copper</td>
<td>10.5 g</td>
<td>3.2 g</td>
<td>7.3 g</td>
</tr>
<tr>
<td>Calcium</td>
<td>8.9 g</td>
<td>2.1 g</td>
<td>6.8 g</td>
</tr>
</tbody>
</table>

9. A mixture of two barium compounds was analysed to determine the dry weight of the mixture. The two barium compounds have the same atomic mass. The results of the tests are given in the table below. List the information in the table to answer the questions which follow.

   a) Name the test used at a) a) and describe how you would use it to determine the dry weight of the mixture. [2 marks]
   b) A barium chloride solution was prepared by dissolving barium chloride in water. How would you determine the mass of water in the mixture? [2 marks]
   c) Write the balanced chemical equation for the reaction of hydrochloric acid with barium carbonate. [2 marks]
   d) Write the balanced chemical equation for the reaction of hydrochloric acid with barium hydroxide. [2 marks]
   e) The mass of dry substances was measured using a balance by accurate weighing. What is the advantage of using a balance? [2 marks]
Working scientifically:
Recording observations

Making and recording observations is an important skill in chemistry. Qualitative observations are what we see and small during reactions. Important types of observations in chemistry and notes on how to record these are shown in the table.

<table>
<thead>
<tr>
<th>Type of Observation</th>
<th>Description</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Colour change</td>
<td>Any change in the colour of the reaction mixture</td>
<td>Solution becomes blue when copper(II) oxide is added to a solution of ammonium hydroxide.</td>
</tr>
<tr>
<td>Odour produced</td>
<td>Any smell produced, such as bubbles in the liquid</td>
<td>When sodium carbonate reacts with an acid, the reaction produces a smell.</td>
</tr>
<tr>
<td>Temperature change</td>
<td>A change in the temperature of the solution</td>
<td>When sodium carbonate reacts with an acid, the temperature of the solution increases.</td>
</tr>
<tr>
<td>Precipitate formation</td>
<td>Any solid that forms as a result of a reaction</td>
<td>When sodium carbonate reacts with an acid, a solid precipitate forms.</td>
</tr>
<tr>
<td>Solubility Check</td>
<td>A check of the solubility of a substance</td>
<td>If a substance is soluble in water, it will dissolve.</td>
</tr>
<tr>
<td>Solubility of solids</td>
<td>A check of the solubility of a solid in water</td>
<td>If a substance is insoluble in water, it will not dissolve.</td>
</tr>
</tbody>
</table>

Questions

1. A laboratory experiment involved adding some magnesium to some copper(II) sulfate solution and recording the observation that the magnesium became covered in copper(II) sulfide. Explain why this happened.
2. State and explain the observations given by the student are correct.
3. Suggest why the student chose to use a filter paper and a spatula to collect the copper(II) sulfate powder and why.
4. Write the balanced chemical equation for this reaction.
5. State the observation made by the student.
6. Copy and complete the table.

<table>
<thead>
<tr>
<th>Type of Observation</th>
<th>Description</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Colour change</td>
<td>Any change in the colour of the solution</td>
<td>The solution changes from blue to yellow when copper(II) oxide is added to a solution of ammonium hydroxide.</td>
</tr>
<tr>
<td>Odour produced</td>
<td>Any smell produced, such as bubbles in the liquid</td>
<td>When sodium carbonate reacts with an acid, the reaction produces a smell.</td>
</tr>
<tr>
<td>Temperature change</td>
<td>A change in the temperature of the solution</td>
<td>When sodium carbonate reacts with an acid, the temperature of the solution increases.</td>
</tr>
<tr>
<td>Precipitate formation</td>
<td>Any solid that forms as a result of a reaction</td>
<td>When sodium carbonate reacts with an acid, a solid precipitate forms.</td>
</tr>
<tr>
<td>Solubility Check</td>
<td>A check of the solubility of a substance</td>
<td>If a substance is soluble in water, it will dissolve.</td>
</tr>
<tr>
<td>Solubility of solids</td>
<td>A check of the solubility of a solid in water</td>
<td>If a substance is insoluble in water, it will not dissolve.</td>
</tr>
</tbody>
</table>
There is air all around us even though we cannot see it. This air contains oxygen which is vital for life, but there was no oxygen in the atmosphere when the Earth was young. This chapter looks at what air is and how the atmosphere has changed over time, it also looks at how the oxygen in the air is used to burn fuels, and how this can produce gases that pollute our atmosphere.

The composition and evolution of the Earth’s atmosphere

The atmosphere today

The Earth is about 4.5 (4000 million) billion years old. For the last 300 million years or so, the composition of the air has been much the same, mainly nitrogen (about 78%) and oxygen (about 21%), with small amounts of noble gases (neon, argon, krypton, and xenon) and carbon dioxide (Figure 9.1). There is also a small amount of water vapour but this varies with weather conditions.

The early atmosphere of the Earth

Scientists believe that the Earth’s atmosphere was very different when it was young. For example, there is evidence that there was little or no oxygen in the atmosphere. However, there is much uncertainty about the Earth when it was young and there are many theories about what the atmosphere was like and how and when it changed.

Mars and Venus have very similar atmospheres to each other. The atmosphere on both planets are mainly carbon dioxide with some nitrogen but with little or no oxygen (Figure 9.2). Given that the Earth is between Venus and Mars, it may have been that the early atmosphere of the Earth was like this. It is thought that the evolution of life on Earth changed our atmosphere, but this did not happen on Venus or Mars.
The composition and evolution of the Earth's atmosphere

One theory about the Earth is that it was very hot when it was formed and there was intense volcanic activity on the planet. Volcanoes release gases from the inside of the planet and the gases in the early atmosphere could have been released from volcanoes in this way (Figure 9.3).

The volcanoes on Earth may have released:
- carbon dioxide ($CO_2$) — which is likely to have been the main gas in the early atmosphere.
- nitrogen ($N_2$) — this may have gradually built up in the atmosphere over time.
- methane ($CH_4$) — probably in small amounts only.
- water vapour ($H_2O$) — this could have condensed and formed the oceans as the Earth cooled down (Figure 9.4).

How the Earth's atmosphere changed

Where did the oxygen come from?

Life evolved on Earth and it is thought that the oxygen in the atmosphere was formed by photosynthesis in living creatures. During photosynthesis, carbon dioxide reacts with water to form glucose and oxygen:

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

Cyanaobacteria were one of the earliest forms of life on Earth (Figure 9.5). They are a type of algae and the first known organisms that photosynthesized. Current fossil evidence is that the oldest cyanobacteria are 2.7 billion years old and it was soon after this that oxygen started to appear in the atmosphere.

Over the next billion years, more complex life forms including plants evolved. As the number of creatures that photosynthesized increased, the amount of oxygen in the air gradually increased and reached a point where animals could evolve.

Where did the carbon dioxide go?

There were two main ways in which carbon dioxide was removed from the atmosphere (Figure 9.6). Much of the carbon dioxide was:
- dissolved in the oceans or
- used in photosynthesis.

Some of the carbon dioxide that dissolved in the oceans reacted to form insoluble carbonate compounds, such as calcium carbonate, that became sediment. Some produced compounds in the oceans that became part of the shells and skeletons of sea creatures. When these creatures died, their shells and skeletons fell into sediment on the sea floor. Over millions of years, all this sediment became sedimentary rocks, such as limestone which is mainly calcium carbonate.

- Figure 9.2 Cyanophyta.
- Figure 9.3 The oceans may have formed the early Earth's atmosphere containing by volatile gases and water.
- Figure 9.4 Cyanobacteria.
- Figure 9.5 Cyanobacteria.
- Figure 9.6 Cyanobacteria.

CO$_2$ in the atmosphere
Greenhouse gases

A lot of carbon dioxide was absorbed by algae and plants for photosynthesis. In the oceans, as algae and other plankton died, their remains were buried in the mud on the sea bed and compacted. Over millions of years this formed crude oil and natural gas that was trapped under rocks. In swamps, the remains of plants were buried and compressed and formed coal, which is a sedimentary rock, over millions of years.

Show you can...
The table shows suggested percentages of atmospheric gases during the early times to present day.

<table>
<thead>
<tr>
<th>Early Times</th>
<th>Present Day</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nitrogen</td>
<td>78.0%</td>
</tr>
<tr>
<td>Oxygen</td>
<td>21.0%</td>
</tr>
<tr>
<td>Carbon Dioxide</td>
<td>0.01%</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>0.00%</td>
</tr>
</tbody>
</table>

Test yourself:
1. Write a balanced equation for the process.
2. What gases make up about 99% of the present atmosphere?
3. What process produces the oxygen in the atmosphere?
4. Write a balanced equation for this process.
5. What gases make up the other 1% of the present atmosphere?
6. How may the oceans have been formed?
7. Where may the gases that formed the Earth’s early atmosphere have come from?
8. What was the early atmosphere made of?
9. Outline how coal was formed.
10. Outline how natural gas and crude oil were formed.
11. What is limestone?

What are greenhouse gases?
The Sun gives off radiation. Some of this reaches the Earth and we call it sunlight. This sunlight contains electromagnetic radiation in the ultraviolet, visible and infrared regions. Much of this radiation passes through the gases in the atmosphere and reaches the surface of the Earth. This provides energy and warmth to the planet.

The Earth also gives off radiation. As the Earth is cooler than the Sun, the radiation given off by the Earth has a longer wavelength than the Sun’s radiation. The radiation given off by the Earth is in the infrared region.

Some of the gases in the atmosphere absorb infrared radiation given off by the Earth but do not absorb the radiation from the Sun. These are known as greenhouse gases and are important for keeping the Earth warm (Figure 8.7). Many occur naturally and include water vapour (H₂O), carbon dioxide (CO₂) and methane (CH₄) (Figure 8.8).

The increase in the amount of greenhouse gases in the atmosphere

The amount of water vapour in the atmosphere varies with weather conditions, and human activities do not have much impact. However, human activities are leading to an increased amount of other greenhouse gases, such as carbon dioxide and methane, in the atmosphere.

The amount of carbon dioxide in the atmosphere

There is a lot of accurate data about the amount of carbon dioxide in the air. It shows clearly that the amount is steadily increasing (Figure 8.9).
One reason why the amount of carbon dioxide is increasing is that very large quantities of fossil fuels are being burned. Carbon dioxide is released when fossil fuels are burned. Figure 9.10 shows how the amount of fossil fuels being burned has increased in recent years. This increase matches the increase in the amount of carbon dioxide in the atmosphere.

A second reason for the increase in carbon dioxide is deforestation (Figure 9.11). Trees remove carbon dioxide from the air as they photosynthesize. In recent years, there has been significant deforestation on the planet as many forests have been cut down but not replaced.

The amount of methane in the atmosphere

There is also data that shows that the amount of methane in the air is increasing (Figure 9.12).

Animal farming produces a lot of methane. Methane is a product of normal digestion in animals. Most of this methane comes from cattle due to the nature of their digestive system (Figure 9.13). It is also produced when manure decomposes. The amount of animal farming has increased in recent years and this is responsible for some of the increase in methane in the atmosphere.

A large amount of waste is buried in the ground in landfill sites (Figure 9.14). Underground the waste decomposes and this also produces a lot of methane gas.

Test yourself

1. What is a greenhouse gas?
2. What is a greenhouse gas that the Earth releases to the atmosphere?
3. What is a greenhouse gas that the Earth absorbs from the atmosphere?
4. What is happening to the amount of carbon dioxide released into the atmosphere?
5. What is happening to the amount of methane released into the atmosphere?
6. What is the reason for this change?
7. What is the reason for this change?
8. What is the reason for this change?
9. What is the reason for this change?
10. What is the reason for this change?

Global warming and climate change

Scientific opinions on global warming

Scientists publish their research in scientific journals. For other scientists and the general public to see (Figure 9.15). Before it can be published, this work is peer-reviewed. This means that it is examined by other scientists who are experts in the same area of science to check that it is scientifically valid.

Based on published peer-reviewed evidence, many scientists believe that the Earth will become warmer due to the human activities that are increasing the amount of greenhouse gases in the atmosphere. They believe that an increase in the temperature at the Earth’s surface, known as global warming, will result in global climate change. However, models that predict the climate are simplifications as there are so many factors that affect the climate and some of these factors are not fully understood. This means that it is very difficult to model the climate fully and there is some uncertainty in this area. For example, there are some scientists who do not believe that the increasing levels of greenhouse gases will make the Earth warmer.

Due to the uncertainty of climate models and significant public interest, there is much speculation and a wide range of opinions are presented in the media on climate change. Some of this may be biased, for example being put forward by industries that use fossil fuels or that promote renewable energy sources. Some may be based on incomplete evidence, perhaps using evidence from one scientific study on its own that does not agree with the majority of published studies.
Sea level rise

As the Earth becomes warmer, sea levels are expected to rise. This may be due to polar ice caps and glaciers melting (Figure 9.15), leading to increased volume of water in the oceans. Another significant factor is that as the water in the oceans becomes warmer, it expands and so has a greater volume. It is difficult to predict by how much sea levels may rise, but there are estimates it could rise by as much as a metre over this century.

Global warming

There is a great deal of evidence from many sources that the temperature of the Earth’s surface has increased in recent years, possibly by about 0.5°C in the last 30 years (Figure 9.16). Most scientists believe that this is largely due to the increased amount of greenhouse gases in the atmosphere.

Climate change and its effects

It is very difficult to predict the effects of the increasing surface temperature of the Earth. Some possible effects are given below.

As sea levels rise, some parts of the world are at risk of flooding (Figure 9.18). For example, there are some Pacific Islands that could be completely submerged. Large areas of countries such as Vietnam, Bangladesh and the Netherlands could be flooded. Some major cities such as London and New York could be at risk in tidal surges if sea levels are higher.

Increased sea levels will also increase coastal erosion. This is where the sea erodes away the rock along the coast. Figures 9.19 and 9.20 show the effect of coastal erosion in Sussex.

Storms

As the Earth warms and the climate changes, there may be more frequent and severe storms (Figure 9.10). However, this is uncertain and there is mixed evidence that this is happening, for example, while there is evidence that there is more rain falling in the most severe storms, there is mixed evidence as to whether the number of hurricanes is increasing or not.

Rainfall

It is expected that global warming will affect the amount, timing and distribution of rainfall. This could increase rainfall in some parts of the world but decrease it in others. For example, there has been more rainfall in northern Europe in recent years but there has been less rainfall in regions near the equator. The number of heavy rainfall events has increased in recent years, but so has the number of droughts, especially near the equator.

Temperature and water stress

As the temperature and climate change, it is expected that this will have an impact on the availability of fresh water. This causes water stress which is a shortage of fresh water and affects all living organisms (Figure 9.21). There have been more droughts in areas near the equator in recent years and this leads to water shortages for humans but also for plants and animals affecting food chains.
Greenhouse gases

Wildlife
Changes in the climate effect wildlife in many ways. For example, plants may flower earlier, birds may lay eggs earlier and animals may come out of hibernation earlier. Some species may migrate further north as temperatures rise. For example, there are species of Dragonfly now being seen in the south of the UK that had only been seen in warmer countries previously. The number of polar bears is expected to fall significantly as the temperature increases as there are fewer suitable places to live (Figure 9.22).

Food production
The weather has a very significant impact on crop production. Changes in the climate may affect the capacity of some regions to produce food due to changes in rainfall patterns, drought, flooding, higher temperatures and the type and number of pests in the region.

Test yourself
10. How do scientists attempt to prove that climate change is the result of human activity?

Carbon footprint
What is a carbon footprint?
The amount of carbon dioxide and other greenhouse gases in the air is increasing and likely to be causing global warming and climate change. Therefore, it is important that we minimise the amount of greenhouse gases released by different activities. A carbon footprint is defined as the amount of carbon dioxide and other greenhouse gases given out over the full life cycle of a product, service or event. For example, the carbon footprint of a plastic bag made from polyethylene could include carbon dioxide released as fuels are burned to provide energy. Next, list:
- a) Polyethylene
- b) Water
- c) Heat energy
- d) Emissions of CO₂
- e) Bacteria
- f) Energy
- g) Water
- h) Energy
- i) Water
- j) Energy

Ways to reduce a carbon footprint
In order to prevent the amount of greenhouse gases in the atmosphere increasing further, it is important to reduce the carbon footprint of products, services and events. Some ways in which this can be done are shown in Table 9.1.

Table 9.1
<table>
<thead>
<tr>
<th>Method of reducing carbon dioxide emissions</th>
<th>Comments</th>
</tr>
</thead>
<tbody>
<tr>
<td>Planting trees and shrubs</td>
<td>This reduces the amount of CO₂ in the atmosphere.</td>
</tr>
<tr>
<td>Increasing energy efficiency in buildings</td>
<td>This reduces the amount of energy needed to heat and cool buildings.</td>
</tr>
<tr>
<td>Using renewable energy sources, such as solar, wind and hydro</td>
<td>This reduces the amount of energy needed to power buildings.</td>
</tr>
<tr>
<td>Using public transport instead of private cars</td>
<td>This reduces the amount of CO₂ in the atmosphere.</td>
</tr>
</tbody>
</table>

Carbon capture and storage
This involves capturing CO₂ from power stations and storing it underground. While this can reduce the amount of CO₂ in the atmosphere, it is not a perfect solution as it is not possible to store all the CO₂ that is produced.

Problems of reducing carbon footprint
There is a need to reduce carbon footprints, but there are some problems that make the reduction in greenhouse gas emissions difficult (Table 9.2).

Figure 9.23: Carbon capture and storage.

Figure 9.24: The carbon in the fuel used for electricity is released into the atmosphere when it is burned.
Table 9.2

<table>
<thead>
<tr>
<th>Description</th>
<th>Suggestion</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gristle digestion</td>
<td>Give this not all agree about the causes and consequences of global climate change. For example, there are some scientific articles to believe that the extra global climate change in the future will contribute to global warming.</td>
</tr>
<tr>
<td>Economic considerations</td>
<td>Many of the cost of reducing carbon footprint are exacerbated. For example, carbon capture and storage will be one of the major contributions to global warming. Due to the high cost of developing countries, it is not easy to implement. One of the major factors is the lack of financial support.</td>
</tr>
<tr>
<td>Domestic international co-operation</td>
<td>The Kyoto Protocol is a major international treaty. Many countries are the signatories of this treaty. However, not all countries are signatories. For example, the United States is not a signatory of this treaty.</td>
</tr>
<tr>
<td>Lack of public information and education</td>
<td>Many people are confused and do not understand the benefits of greenhouse gases and climate change. There is a need to educate the public about these issues.</td>
</tr>
<tr>
<td>Adverse changes</td>
<td>The world's energy consumption is rapidly increasing. There's a lot of room for reduction in the immediate future. The United States is a major emitter of greenhouse gases. In recent years, the United States has made significant progress in reducing its greenhouse gas emissions.</td>
</tr>
</tbody>
</table>

Figure 9.23: Greenhouse gas (GHG) emissions in the US 1990–2013

Test yourself
1. What does your carbon footprint mean?
   a. It is the amount of wastewater a person uses in a day.
   b. It is the amount of greenhouse gas emissions a person produces.

2. What is a way to reduce your carbon footprint?
   a. Use more alternative energy sources.
   b. Decrease the amount of meat you eat.

Common atmospheric pollutants and their sources

What is formed when fuels burn?
Most fuels, including coal, oil, and gas, contain carbon and/or hydrogen and may also contain some sulfur.

When fuels burn, they are oxidized which means that they react with oxygen and the atoms in their molecules combine with the oxygen. For example, when fuels containing hydrogen burn, the hydrogen combines with oxygen to form water vapor. When fuels containing sulfur burn, the sulfur combines with oxygen to form sulfur dioxide.

For fuels containing carbon, complete combustion takes place in a process called carbon dioxide combustion. In a proper supply of oxygen, complete combustion takes place which gives carbon dioxide and water vapor (a form of carbon dioxide).

Table 9.3 gives some equations which are examples of reactions taking place when fuels burn.
Pollution from burning fuels

The burning of fuels is a major source of pollutants in the atmosphere. Table 9.5 shows some of these, how they are formed and how the problems they cause could be reduced.

Table 9.2

<table>
<thead>
<tr>
<th>Pollution in fuel</th>
<th>Medium</th>
<th>Fresh combustion</th>
<th>Combustion reactions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methane (CH₄)</td>
<td>Methane is a hydrocarbon fuel</td>
<td>CH₄ + O₂ → CO₂ + H₂O</td>
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<td>Methane → water + carbon dioxide</td>
</tr>
</tbody>
</table>

Figure 9.11: Tower Bridge has been cleared to remove smuts caused by air pollution.
Test yourself

23. Write both a word equation and a balanced equation to show what happens when each of the following substances undergoes complete combustion:
   a. Hydrogen, H₂
   b. Carbon monoxide, CO
   c. Sulfur, S
   d. Potassium, K
   e. Iron, Fe
   f. Phosphorus, P

24. A liquid is important fuel used in cars. It is
   a. One of the chemicals in petrol is octane, C₈H₁₈. Under what conditions are the following gases formed when substances burn?
   b. Water vapor (H₂O)
   c. Carbon dioxide (CO₂)
   d. Carbon monoxide (CO)

25. Explain why some sulfur dioxide may be formed when sulfur burns.

26. What two problems do the pollutants sulfur dioxide and nitrogen oxides cause in the air?
   a. Some untreated wastewater can be given off in exhaust fumes from cars. Why might this be a problem?
   b. Carbon monoxide can be formed in the incomplete combustion of methane. CH₄.
      i. What is incomplete combustion?
      ii. Why is carbon monoxide toxic?
      iii. Why is it difficult to detect carbon monoxide?
      iv. Why can it be formed on the incomplete combustion of methane (CH₄)?

27. Sulfur can be formed in the incomplete combustion of methane (CH₄).
   a. Explain why sulfur can be formed.
   b. What problems can it cause?

Chapter review questions

1. Copy and complete the gaps in the paragraph below.

   The earth is thought to be about _______ years old. One theory is that there was a lot of activity when the earth was young and this gave it the same atmosphere.
   a. The air today consists of 78% nitrogen, 21% oxygen and 1% other gases. It received its air through
      i. The present-day atmosphere is the same as when the earth was young.
   b. Some scientists believe that the oceans were formed from water vapor given off by volcanic eruptions. How could the oceans have formed from this water vapor?
   c. Scientists believe that the air that the earth is thought to have had when it was young was carbon dioxide; the evolution of life is thought to have changed the earth's atmosphere when animals and plants evolved that photosynthesized.
   d. Which two gases in the solar system are thought to be mainly carbon dioxide?
   e. Write a word equation and a balanced equation for photosynthesis.

2. Copy and complete the gaps in the paragraph below.

   a. Sulfur dioxide is formed when sulfur is burned in air. It causes serious injury to the respiratory system.
   b. Nitrogen dioxide (NO₂) is formed when nitrogen is heated in air. It causes serious injury to the respiratory system.
   c. Carbon monoxide is formed when carbon is burned in air. It causes serious injury to the respiratory system.
   d. Carbon dioxide (CO₂) is formed when carbon is burned in air. It causes serious injury to the respiratory system.
   e. Water vapor (H₂O) is formed when water is heated in air. It causes serious injury to the respiratory system.

3. Which of the following gases is formed from the combustion of methane (CH₄)?
   a. Carbon dioxide (CO₂)
   b. Carbon monoxide (CO)
   c. Sulfur dioxide (SO₂)
   d. Nitrogen (N₂)
   e. Water vapor (H₂O)
a) Methane (CH₄) is a greenhouse gas. Describe the ways in which human activities are increasing the amount of methane in the atmosphere.

b) Carbon dioxide (CO₂) is a greenhouse gas. Describe two ways in which human activities are increasing the amount of carbon dioxide in the atmosphere.

c) There is a lot of interest taking place to reduce the carbon footprint of many processes.

10. Write a balanced equation for the complete combustion of each of the following fuels.

a) Propane, C₃H₈
b) Methane, CH₄

c) Methanol, C₂H₅OH

d) There is a small amount of nitric oxide (NO) found in the process of making car exhaust. Write a balanced equation for the complete combustion of nitric oxide.

11. The Sun's radiation is in the ultraviolet, visible, and near infrared region of the electromagnetic spectrum. The radiation given out by the Earth is in the infrared region.

12. a) How does the wavelength of the Earth's radiation compare to the Sun's?

b) Why are carbon dioxide, methane, and water vapour greenhouse gases?

Practice questions

1. Which one of the following makes up about 2% of the Earth's atmosphere?

- Carbon dioxide
- Nitrogen
- Oxygen

2. A company has decided to reduce the amount of sulfur dioxide from the waste gases which it emits to its factories. Which of the following substances in solution would most effectively remove the sulfur dioxide from the waste gases?

- Calcium hydroxide
- Sodium chloride
- Sulphuric acid

3. Mars is often called the red planet due to the presence of hematite, which contains iron (II) oxide. What is hematite?

- A rust-like substance
- A mineral
- A chemical compound

4. The table shows some information about different fossil fuels.

<table>
<thead>
<tr>
<th>Fossil fuel</th>
<th>Appearance</th>
<th>Carbon</th>
<th>Hydrogen</th>
<th>Sulfur</th>
<th>Nitrogen</th>
</tr>
</thead>
<tbody>
<tr>
<td>Coal</td>
<td>Charcoal</td>
<td>75</td>
<td>8</td>
<td>0.5</td>
<td>2</td>
</tr>
<tr>
<td>Petroleum</td>
<td>Liquid</td>
<td>85</td>
<td>6</td>
<td>1.5</td>
<td>1.5</td>
</tr>
<tr>
<td>Natural gas</td>
<td>Gas</td>
<td>80</td>
<td>8</td>
<td>1.5</td>
<td>2</td>
</tr>
<tr>
<td>Fried oil</td>
<td>Liquid</td>
<td>75</td>
<td>2</td>
<td>2.5</td>
<td>1.5</td>
</tr>
<tr>
<td>Lignite</td>
<td>Solid</td>
<td>55</td>
<td>20</td>
<td>2.5</td>
<td>1.5</td>
</tr>
</tbody>
</table>

5. a) Write the CuSO₄ in water.

b) Iron(II) oxide.

6. a) Write a balanced equation for the thermal decomposition of K₂CO₃.

b) Write a balanced equation for the thermal decomposition of Na₂CO₃.

7. a) Write a balanced equation for the thermal decomposition of Na₂CO₃.

b) Write a balanced equation for the thermal decomposition of K₂CO₃.

8. a) Write the balanced equation for the thermal decomposition of Na₂CO₃.

b) Write the balanced equation for the thermal decomposition of K₂CO₃.

9. a) Write the balanced equation for the thermal decomposition of Na₂CO₃.

b) Write the balanced equation for the thermal decomposition of K₂CO₃.

10. a) Write the balanced equation for the thermal decomposition of Na₂CO₃.

b) Write the balanced equation for the thermal decomposition of K₂CO₃.

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b) Write the balanced equation for the thermal decomposition of K₂CO₃.

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b) Write the balanced equation for the thermal decomposition of K₂CO₃.

13. a) Write the balanced equation for the thermal decomposition of Na₂CO₃.

b) Write the balanced equation for the thermal decomposition of K₂CO₃.

14. a) Write the balanced equation for the thermal decomposition of Na₂CO₃.

b) Write the balanced equation for the thermal decomposition of K₂CO₃.

15. a) Write the balanced equation for the thermal decomposition of Na₂CO₃.

b) Write the balanced equation for the thermal decomposition of K₂CO₃.

16. a) Write the balanced equation for the thermal decomposition of Na₂CO₃.

b) Write the balanced equation for the thermal decomposition of K₂CO₃.

17. a) Write the balanced equation for the thermal decomposition of Na₂CO₃.

b) Write the balanced equation for the thermal decomposition of K₂CO₃.

18. a) Write the balanced equation for the thermal decomposition of Na₂CO₃.

b) Write the balanced equation for the thermal decomposition of K₂CO₃.

19. a) Write the balanced equation for the thermal decomposition of Na₂CO₃.

b) Write the balanced equation for the thermal decomposition of K₂CO₃.

20. a) Write the balanced equation for the thermal decomposition of Na₂CO₃.

b) Write the balanced equation for the thermal decomposition of K₂CO₃.
1. Scientists in the Antarctic have measured the concentration of carbon dioxide in air bubbles in the ice there, and this has allowed them to estimate the atmospheric concentration of carbon dioxide over many thousands of years, as shown in graph 1. Graph 2 shows the change in average temperature of the Earth over the same period. Graphs 1 and 2 show the relationship between the Earth's temperature and the concentration of carbon dioxide in the atmosphere. Do the graphs show that an increase in atmospheric carbon dioxide concentration increases global warming? Support your answer. (2 marks)

2. Scientists are developing methods to estimate climate change. State two methods which have been used to estimate climate change by removing carbon dioxide. (3 marks)

3. Governments have set targets to reduce emissions of carbon dioxide. One way of doing this is to increase carbon dioxide concentrations. Describe how carbon dioxide emissions could be used to reduce carbon dioxide emissions. (3 marks)

Working scientifically: Communicating scientific conclusions

When drawing conclusions from data about burning fossil fuels and air pollution, it is important to realize that a number of different influences can affect the data and these need to be taken into consideration. There are many things which a scientist must do to produce valid results and conclusions.

- A scientist may take a set of measurements or make some observations and draw conclusions from them. However, to ensure valid conclusions are drawn, it is important that:
  1. The experiment or observation is carefully controlled, meaning that the results are reproducible. This increases the reliability of the conclusion.
  2. Independent results are published, whether they are for or against the hypothesis. This is known as peer review.
  3. Scientists publish their results in journals, which are then evaluated by other experts in the same field.

Cold fusion

Nuclear fusion in a test tube developed by Irvin Tovch — Russian Times

Scientists pursue endless power source — The Times

Scientists claim techniques to control nuclear fusion — Russian Times

If you had read the newspaper headlines shown above in March 1989, you might have believed that the world's energy generation problems could be solved overnight! A team of scientists announced that they had successfully induced nuclear fusion to happen using simple laboratory equipment operating at relatively low temperatures. This became known as “cold fusion.”

Nuclear fusion involves two atomic nuclei overcoming the repulsion between them and merging together to make a larger nucleus. It is only known to happen at very high temperatures. This released tremendous amounts of energy and was the energy source of the stars. In 1989, reports that fusion was occurring at relatively low temperatures generated a great deal of attention and raised hopes of a cheap and abundant source of energy.

Pros and cons were not the concern at the start of this scientific work. They announced their findings at a press conference before...
Questions

1. What was unusual about the way in which Hirsch and Nappi assessed their results?
2. What procedures should a scientist follow before presenting their findings?
3. A concentration of particulates in the air at a town centre was measured over several days. The number of patients seeking medical treatment for asthma was recorded over the same days. The results were plotted in the graph shown.

   ![Graph of concentration of particulates in the air over time]

4. What correlation does the graph show?
5. Evaluate the validity of the data.
6. The results of this experiment were presented in a newspaper. The headline claims: 

   **Asthma is caused by particulates in the air**

   a. How much confidence can be placed in the newspaper claim?
   b. The particulates that pollute the air are made of carbon. Explain how carbon particulates get into the air.

7. The chapter covers specification points 4.10.1 to 4.10.6 and is called Using the Earth’s resources. It covers using the Earth’s resources and obtaining possible water, life cycle assessment and recycling, using materials, the Haber process and the use of NPK fertilizers.
Previously you could have learned:

**Useful products are made from the raw materials found on the Earth.**

Metals are extracted from ores found in rocks; the method used to extract the metal depends on the reactivity of the metal.

- Alloys are mixtures of metals with small amounts of other metals or carbon; alloys are harder than pure metals.
- Polymers (plastics) are long-chain molecules made from joining lots of short molecules together.
- Some chemical reactions are reversible and can reach a state of equilibrium in a closed system.

Test yourself on prior knowledge:

1. What raw materials do we use to make each of the following?
   - Metals
   - Water
   - Oxygen
   - Cotton
2. Some metals are extracted from ores by heating with carbon and some are extracted by electrolysis.
   - Identify two metals that could be extracted from their ores by heating with carbon.
   - Identify two metals that could be extracted from their ores by electrolysis but not by heating with carbon.
3. What is an alloy?
4. Give one property of an alloy that is different from the metal it is made of.
5. What is a polymer?
6. Name two polymers.
7. Some reactions are reversible. What does this mean?

Using the Earth's resources:

- **The Earth’s resources**
  - Everything that we need for life is provided by the Earth’s resources.
  - These resources include:
    - Rocks in the ground
    - Fuels such as coal, oil and natural gas found underground
    - Plants and animals (productive) and agriculture
    - Fresh water and sea water
    - Air
    - Sunlight
    - Wind
  - Table 10.1 shows some examples of how we use these resources to provide energy and essential substances for everyday life.

<table>
<thead>
<tr>
<th>Table 10.1</th>
<th>Energy</th>
<th>Water</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sources</td>
<td>Usage</td>
<td></td>
</tr>
<tr>
<td>Oil</td>
<td>Fuel</td>
<td></td>
</tr>
<tr>
<td>Natural gas</td>
<td>Heating</td>
<td>Domestic</td>
</tr>
<tr>
<td>Coal</td>
<td>Heating</td>
<td>Industrial</td>
</tr>
<tr>
<td>Solar panels</td>
<td>Heating</td>
<td>Renewable energy</td>
</tr>
</tbody>
</table>

- **Sustainable development**
  - Many of the Earth’s resources are finite. This means that we cannot replace them once we have used them. For example, supplies of crude oil and many metal ores are limited and if we continue to use them as we do now we will run out. In contrast, this, some resources are renewable, which means we can replace them once we have used them. Instead, which are fuels made from plants, are good examples of renewable resources. Examples of biofuels include ethanol and biodiesel that can be made from crops and so can be replaced once used.
  - Sustainable development is where we use resources to meet the needs of people today without preventing people in the future from meeting their needs. Table 10.2 shows some ways in which we can meet our needs today in more sustainable ways.

<table>
<thead>
<tr>
<th>Table 10.2</th>
<th>Sustainable energy</th>
<th>Sustainable transportation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Wind energy</td>
<td>No fossil fuel</td>
<td>Biofuel replace petrol</td>
</tr>
<tr>
<td>Solar power</td>
<td>No fossil fuel</td>
<td>Biofuel replace petrol</td>
</tr>
<tr>
<td>Hydro power</td>
<td>No fossil fuel</td>
<td>Biofuel replace petrol</td>
</tr>
<tr>
<td>Biomass</td>
<td>No fossil fuel</td>
<td>Biofuel replace petrol</td>
</tr>
<tr>
<td>Geothermal</td>
<td>No fossil fuel</td>
<td>Biofuel replace petrol</td>
</tr>
</tbody>
</table>

- **Key terms**
  - Finite resources: A resource that cannot be replaced once it is used.
  - Renewable resources: A resource that we can replace once we have used it.
  - Sustainable development: Using resources to meet the needs of people today without preventing people in the future from meeting their needs.

- **Figure 10.1** Using the wind to generate electricity is sustainable.

- **Figure 10.2** This shows how a wind turbine works.
1. Which of the Earth’s natural resources are used to make the following substances?
   a) metals  
   b) plastic  
   c) cloth  
   d) water  
   e) forests
2. What is a finite resource?
3. What is a renewable resource?
4. Name two finite resources.
5. Name two renewable resources.
6. a) What is sustainable development?  
    b) Why is it so important?
7. a) The use of natural gas as a fuel for electricity generation is not sustainable. Suggest how we could generate electricity in a sustainable way.
8. Which are the uses of water?  
   a) Food production  
   b) Energy production  
   c) Industry  
   d) Domestic use
9. a) What is pollution?  
    b) What is a pollutant?
10. Materials for some uses have to be of very high quality. For example, glass used for making jars for cooking and metals used for making tools must be of very high quality to have the required properties. For other uses, such as plastic for making rulers, the quality of the material is less important. In general, when using recycled products to produce materials, the better the separation of the materials when they are recycled, the higher the quality of the final material produced.

When metals, glass or plastics are recycled, the different types are separated. For example:
- Iron, steel, aluminium and copper are common metals that are separated from other metals.
- Coloured, brown and green glass are separated.
- Plastics such as high-density poly(ethylene), low-density poly(ethylene), PET and PVC are separated (Figure 10.3).

A good example of recycling is the use of scrap steel to make new steel. (Figure 10.4). Steel is made using mainly iron which is extracted from iron ore in a blast furnace. By adding scrap steel to the same furnace, the amount of iron that needs to be extracted to make steel is reduced.

**Life cycle assessment**

LCA is carried out to assess the impact of a product on the environment throughout its life. This includes the extraction of raw materials, its manufacture, its use and its disposal at the end of its useful life. Factors that are considered include the use of land and sustainability of raw materials (including those used for the packaging of the product), the use of energy at all stages and the production and disposal of waste products (including pollutants) at all stages. Transportation and distribution are included at all stages.

Simple life cycle assessments of the use of plastic and paper to make shopping bags are shown in Table 10.4 for comparison.
In a life cycle assessment, the use of raw materials, energy and water plus the production of some waste can be quite easy to quantify. However, it can be difficult to quantify the effects of some pollutants. For example, it is hard to judge how much damage to the environment is caused by the release of the greenhouse gas methane from the rotting of paper bags in landfills. This means that a life cycle assessment involves some personal judgements to make a valid judgement to decide whether use of a product is good or not. This means that the process is not completely objective and so judgements about the benefits or harm of the use of a product may be a matter of opinion.

Life cycle assessments can be produced that do not show all aspects of a product's manufacture, use and disposal. These could be inflated, for example by the manufacture of a product to support the use of their product.

### Show you can...

#### The uses of water

#### Producing potable water

**Types of water**

Water is essential for life and humans need water that is safe to drink (Figure 10.5). Water that is safe to drink is called potable water. Potable water is not pure water as it contains some dissolved substances, but these are at low levels that are safe. There should also be safe levels of minerals in potable water and ideally none.

Potable water and other types of water are described in Table 10.5.

#### Water treatment

In the UK, we produce our potable water through a fresh-water treatment process. Some parts of the UK, fresh water comes mainly from rivers and reservoirs, but in other areas, much water comes from ground water sources. There are many stages in making this water safe to drink, but two of the main stages are shown in Table 10.6.

---

**Table 10.4**

<table>
<thead>
<tr>
<th>Manufacturing</th>
<th>Paper recycling</th>
<th>Free shipping bags</th>
</tr>
</thead>
<tbody>
<tr>
<td>Source</td>
<td>12</td>
<td>2.2</td>
</tr>
<tr>
<td>Sustainability</td>
<td>Yes</td>
<td>Brakes are burnt as a by-product.</td>
</tr>
<tr>
<td>Weight</td>
<td>1.5</td>
<td></td>
</tr>
<tr>
<td>Number of items</td>
<td>2</td>
<td></td>
</tr>
<tr>
<td>Cost</td>
<td>$1.50</td>
<td></td>
</tr>
</tbody>
</table>

**Table 10.5**

<table>
<thead>
<tr>
<th>Type of water</th>
<th>Description</th>
<th>Characteristics</th>
<th>Advantages</th>
<th>Disadvantages</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fresh water</td>
<td>Water that contains only water molecules and nothing else</td>
<td></td>
<td>High</td>
<td></td>
</tr>
<tr>
<td>Potable water</td>
<td>Water that is safe to drink</td>
<td>Yes</td>
<td>Yes</td>
<td></td>
</tr>
<tr>
<td>Source water</td>
<td>Water found in rivers, lakes and underground rocks and minerals</td>
<td>Yes</td>
<td>Yes</td>
<td></td>
</tr>
<tr>
<td>Desalinated water</td>
<td>Water in the process of being filtered</td>
<td>Yes</td>
<td>Yes</td>
<td></td>
</tr>
<tr>
<td>Groundwater</td>
<td>Water from rivers, lakes, and underground rocks and minerals</td>
<td>Yes</td>
<td>Yes</td>
<td></td>
</tr>
</tbody>
</table>

**Table 10.6**

<table>
<thead>
<tr>
<th>Stage</th>
<th>Water</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>Purification</td>
<td>Water that is passed through filters and is clean of most of the water and debris.</td>
<td></td>
</tr>
<tr>
<td>Disinfection</td>
<td>Water that is passed through a disinfectant and is clean of most of the water and debris.</td>
<td></td>
</tr>
</tbody>
</table>

---

**Figure 10.5 Harnatt needs water that is safe to drink.**
In some parts of the world, such as Saudi Arabia, the United Arab Emirates (UAE) and parts of Syria, there is little fresh water but there is lots of sea water. Potable water can be made from sea water by desalination. Two methods of doing this are distillation and reverse osmosis (Table 10.7).

Both methods require a lot of energy. In distillation the energy is needed to boil the water. In reverse osmosis the energy is needed to create the pressure. Reverse osmosis is becoming more popular in places with less sea water but little fresh water as it is cheaper than distillation. However, due to the high energy costs both methods are more expensive than producing potable water from fresh water.

### Analysis and purification of water samples

Different experimental techniques can be used to study the composition of water. The type of equipment and chemical reagents used will depend on the nature of the problem being studied. Before using a new technique, the operator should familiarize herself/himself with the procedure involved and the hazards associated with its use. The less suitable technique will be used. Pure water samples are placed in a beaker. The solution is heated to just below boiling point; the pressure inside the beaker is reduced, the gas burns, and the vapour is condensed to water. The water is then collected. The more the boiling point changes from the expected boiling point of water, the less suitable the technique is.

Water has been purified by filtration to form drinking water suitable for at least 40% of the world’s population. The filtration apparatus used to purify water from different sources, such as well water, is shown in diagram 1. To determine the identity of dissolved ions in water, the ions studied in Chapter 11 will be used.

#### Questions

1. Name the vessel in each solution.
2. Name the beaker in each solution.
3. What is a distillation apparatus used for.
4. What is the purpose of the filter paper in the apparatus.
5. What is the purpose of the test-tube in the apparatus.
6. What is the purpose of the thermometer in the apparatus.
7. State and explain what happens to the distilled water in the solution during the process.
8. Distillation involves two processes: evaporation and condensation. What is the purpose of each process.
9. What is the purpose of the condenser in the apparatus.
10. What is the purpose of the condenser in the apparatus.
11. Name the pieces of apparatus in diagram 1.
12. Describe how you would test the purity of the distilled water.
13. State why drinking water is not suitable for the environment.
14. How would you determine the purity of the water before and after distillation.
15. How would you analyse the sample of mineral water contained dissolved solids?

### Waste water treatment

Large amounts of waste water are continuously produced. This waste water comes from domestic, industrial, and agricultural uses of water and has to be treated before the water can safely be returned to the environment. The waste water is collected in underground pipes and taken for treatment. In some places, all waste water is collected together in combined sewers, but in other places there are separate systems for different types of waste water. These separate systems are more efficient as different types of waste water require different treatment.
Sewage is an example of waste water and includes water used to flush toilets that contains human waste and toilet paper. It also includes water from sinks, baths, and showers. Sewage needs to be treated to remove any objects in the water, the organic matter (the human waste) and harmful microorganisms. Waste water from agricultural and industrial sources also needs treating as it may also contain organic matter, harmful microorganisms and some harmful chemicals.

Key stages in the treatment of waste water include (Figure 10.6):
- Screening and grit removal - this removes large solids from the waste water.
- Sedimentation - this separates the human waste from the rest of the water which is called effluent.
- Anaerobic treatment of effluent (from sedimentation) - all the waste is broken down through the effluent in aeration tanks which leads to good bacteria killing harmful bacteria.
- Anaerobic treatment of sludge (from sedimentation) - in the absence of all the bacteria produce methane from sludge.

Test yourself:
18. What is sewage water?
19. What is the role of sewage water production from human waste? What is its role?
20. A key stage in the production of sewage water is filtration. Why and how is this done?
21. What is the role of sewage water in the production process of sewage water?
22. In this stage (in the production process of sewage water), how does the water change in which of these ways?
23. What is the role of sewage water?
24. More sewage water is produced by desalination in Saudi Arabia than in any other country. Why is water produced this way in Saudi Arabia?
25. Describe how sewage water is produced from sea water by desalination.
26. Describe how sewage water is produced from sea water made by reverse osmosis.
27. Give three sources of sewage water.
28. For each of the following stages of sewage water production, what happens?
   - screening
   - flotation
   - aeration
   - anaerobic treatment of effluent
   - anaerobic treatment of sludge

Show you can:

<table>
<thead>
<tr>
<th>Method of producing water</th>
<th>Waste water</th>
<th>Ground water</th>
<th>Soft water</th>
</tr>
</thead>
<tbody>
<tr>
<td>Desalination</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Groundwater</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Reverse Osmosis</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

In 2004 the Israeli desalination plant, the first and only desalination plant in the UAE, opened and is able to produce 1% of the UAE’s water.

Metals and other materials

Metals
Metals are very important materials that have many uses. For example:
- cars, trains and planes are made from metal
- large buildings have metal frameworks
- all electrical devices have metal circuits and wires
- tools are made from metal.

Metals are extracted from compounds found in rocks (see Chapter 4).

The metal compounds found in rocks are called ores. An ore is a rock from which a metal can be extracted for profit. There are several ways to extract the metals from rocks but the method used depends on the reactivity of the metal.
Alloys

Pure metals such as aluminium, copper, gold and iron are too soft for many uses. They can be made harder and more useful by turning them into alloys. An alloy is a mixture of a metal with small amounts of other metals or carbon.

Pure metals are soft because all the atoms are the same size and so layers of atoms can slide over each other. Alloys are harder because there are some different sized atoms present which makes it much more difficult for atoms to slide (Figure 10.7).

Some examples of alloys are described in Table 10.8.

Table 10.8

<table>
<thead>
<tr>
<th>Metal</th>
<th>Other elements</th>
<th>Hardened?</th>
<th>Uses</th>
</tr>
</thead>
<tbody>
<tr>
<td>Iron</td>
<td>Carbon, silicon</td>
<td>Yes</td>
<td>Steel, tools, machinery</td>
</tr>
<tr>
<td>Brass</td>
<td>Copper, zinc</td>
<td>Yes</td>
<td>Bronze, plumbing fixtures</td>
</tr>
<tr>
<td>Bronze</td>
<td>Copper, tin</td>
<td>Yes</td>
<td>Coins, ornamental metal</td>
</tr>
<tr>
<td>Steel</td>
<td>Carbon</td>
<td>Yes</td>
<td>Automotive, construction</td>
</tr>
<tr>
<td>Gold</td>
<td>Silver, copper</td>
<td>No</td>
<td>Fine jewelry, dental work</td>
</tr>
<tr>
<td>Silver</td>
<td>Copper, tin</td>
<td>No</td>
<td>Coins, silverware</td>
</tr>
</tbody>
</table>

Gold alloys

The purity of gold alloys is measured in carats. Pure gold is 24 carat. An alloy containing 75% gold would be 18 carat. Jewellery is not made from pure gold because it is too soft and would lose its shape and any engravings (Figure 10.8).

Example

An alloy of gold is found to be 18 carat. Calculate the percentage of gold in this alloy.

Answer

If gold is 18 carat, then it contains 18/24 of the total mass. This is 0.75, or 75% gold. There are also many other alloys that contain gold.

Example

An alloy of gold is found to contain 99% gold. Calculate the number of carats this alloy has.

Answer

If gold is 99% pure, then it contains 0.99 of the total mass. This is 99 carats. Therefore, this alloy contains 99 carats.

Steels

Steels are alloys of iron mixed with small amounts of carbon and/or other metals. Steels are used more than other alloys due to their low cost and high strength.

Corrosion of metals

Corrosion is the destruction of materials by chemical reactions with substances in the environment. Many metals corrode when they come into contact with oxygen and/or water.

What is rusting?

Rusting occurs when iron and steel are exposed to oxygen and water. The iron reacts with oxygen and water to form iron oxide, which is rust. Rust is a dark brownish-red substance that forms on the surface of iron and steel when they are exposed to oxygen and water.

Figure 10.11: A patch of rust on a car.
When iron (or the iron in steel) corrodes, it reacts with oxygen and water to produce rust.

\[ \text{Iron} + \text{water} + \text{oxygen} \rightarrow \text{rust} \]

This can be shown by experiment where nails are placed in a series of boiling tubes under different conditions (Table 10.10).

<table>
<thead>
<tr>
<th>Tube</th>
<th>Reaction</th>
<th>Nails present</th>
<th>Nails present</th>
<th>Result</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>air-pump</td>
<td>yes</td>
<td>no</td>
<td>red rust</td>
</tr>
<tr>
<td>2</td>
<td>water</td>
<td>no</td>
<td>yes</td>
<td>black rust</td>
</tr>
<tr>
<td>3</td>
<td>water</td>
<td>yes</td>
<td>yes</td>
<td>no rust</td>
</tr>
</tbody>
</table>

The nail does not rust in Tube 1 where there is no water present. It also does not rust in Tube 2 where there is no air present. However, when there is both air and water present the nail rusts. It can also be shown by experiment that it is the oxygen in the air that reacts with the nail rusts in the presence of air and water.

Preventing corrosion

Two basic ways to prevent the corrosion of metals are to put a coating on the surface of the metal to act as a protective layer or to use sacrificial protection where a more reactive metal is attached to the metal (Table 10.11).

Table 10.11

<table>
<thead>
<tr>
<th>Coating</th>
<th>How it works</th>
<th>Example 1</th>
<th>Example 2</th>
</tr>
</thead>
<tbody>
<tr>
<td>Surface coating</td>
<td>It acts as a barrier to prevent the metal from coming into contact with the \textbf{air} and \textbf{water}</td>
<td>Zinc coating on a car</td>
<td>Paint coating on a bicycle frame</td>
</tr>
<tr>
<td>Sacrificial protection</td>
<td>The metal reacts with corrosion products and forms a protective layer</td>
<td>Steel with a layer of zinc</td>
<td>Copper with a layer of zinc</td>
</tr>
</tbody>
</table>

Steel can be galvanised, which combines both of these methods. (Figure 10.13). The steel is coated in a layer of zinc. Surface coatings often stop working if the surface is scratched and the metal exposed. However, if the layer of zinc is scratched the steel is still protected by sacrificial protection as the zinc is more reactive.

Test yourself

19 a. What is corrosion?
   b. What is rusting?
   c. Write a word equation to show what happens when iron rusts.
   d. Write a word equation to show what happens when aluminium rusts.
   e. Explain how each of the following prevents steel from rusting.
      i. painting the steel
      ii. coating the steel with a layer of chromium by electroplating
      iii. galvanising the steel
   f. Explain why aluminium does not appear to corrode.

Show you can...

A pipe can be protected from corrosion by sacrificial protection. All parts of the pipe are made of iron. The iron is coated with a layer of zinc which is then coated with a layer of paint.

Extraction of copper

The Earth’s reserves of copper ore are limited. We have already reached a point with copper metal where there are only low-grade ores remaining. Low-grade ores are ones which only contain a small percentage of metal compounds. Due to this, scientists are having to develop new methods to extract copper.
One way of doing this is to use **physioleaching** to extract copper from ore that contains copper compounds. (Figure 33.14.) Some plants are very good at absorbing copper compounds from the ground through their roots. These plants can be grown in land containing copper compounds. These plants can then be burned leaving ash that is rich in copper compounds.

Copper can be extracted from this plant ash or from smelting low-grade ore using **physioleaching** or **leaching** of the copper compound with a weak acid such as sulfuric acid. In each case, insoluble copper compounds need to produce a solution containing soluble copper compounds. This solution is known as a **leachate**. Some bacteria are very efficient at this and so biolatching can be very effective.

The copper can be extracted from the solution containing copper compounds. This can be done by electrolysis or by using soap in a displacement reaction (Table 33.12).

Traditional methods of metal extraction involve mining, with digging, moving and disposing of huge amounts of soil. These new methods for copper extraction are very different and do not involve this and so are better for the environment.

<table>
<thead>
<tr>
<th>Test yourself</th>
</tr>
</thead>
<tbody>
<tr>
<td>23. What is a low-grade ore?</td>
</tr>
<tr>
<td>24. Copper can be extracted from low-grade ores by biolatching. What is biolatching?</td>
</tr>
<tr>
<td>25. Leachate solutions containing copper compounds can be produced by physioleaching or leaching. Describe a procedure for producing leachate solutions containing copper compounds.</td>
</tr>
<tr>
<td>26. Two ways to extract copper metal from leachate solutions are by electrolysis and by displacement using soap. Explain why electrolysis is not used.</td>
</tr>
<tr>
<td>27. In the displacement process:</td>
</tr>
<tr>
<td>a) What is the reaction?</td>
</tr>
<tr>
<td>b) Explain the reaction for the formation of copper in this process.</td>
</tr>
<tr>
<td>28. Show you can...</td>
</tr>
<tr>
<td>Extracting copper metal from leachate solutions in laboratories and by investigating one of the processes when used on a industrial scale.</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Other materials</th>
</tr>
</thead>
<tbody>
<tr>
<td>Polymers, glass, ceramics and composites are important materials made by chemical reactions from the Earth’s natural resources.</td>
</tr>
</tbody>
</table>

**Polymers**

Polymers are long chain molecules made from joining lots of short molecules called monomers together. Examples of addition polymers include polyethylene, polypropylene, nylon, PVC and Teflon. Examples of condensation polymers include polyesters and polyamides. Different polymers have different properties and so have different uses. The properties of a polymer depend on:

- which monomer(s) they are made from, and
- the conditions under which they are made.

Table 33.13 shows how using different monomers produces different polymers.
Table 10.13. Some polymers and their uses.

<table>
<thead>
<tr>
<th>Property</th>
<th>Formula</th>
<th>Polymer Chain</th>
<th>Uses of polymer</th>
</tr>
</thead>
<tbody>
<tr>
<td>Epoxy</td>
<td>HO(CH₂)₃OH</td>
<td>Polyetheramide</td>
<td>Coating, adhesives</td>
</tr>
<tr>
<td>Acrylic</td>
<td>CH₂=CH-CO₂H</td>
<td>Acrylonitrile</td>
<td>Polymers, adhesives, coatings</td>
</tr>
<tr>
<td>Polyethylene</td>
<td>(CH₂)₁₀</td>
<td>Polyethylene</td>
<td>Packaging, films, fibers</td>
</tr>
<tr>
<td>Polypropylene</td>
<td>(CH₂)₉-CH₂</td>
<td>Polypropylene</td>
<td>Plastics, fibers, films</td>
</tr>
</tbody>
</table>

However, polymers with different properties can be made using the same monomer, but by changing the reaction conditions. For example, low-density (LD) and high-density (HD) polyethylene are both made from the monomer ethylene but made under different conditions (Table 10.14).


<table>
<thead>
<tr>
<th>Property</th>
<th>Low-density (LD) polyethylene</th>
<th>High-density (HD) polyethylene</th>
</tr>
</thead>
<tbody>
<tr>
<td>Density</td>
<td>0.92 g/cm³</td>
<td>0.94 g/cm³</td>
</tr>
<tr>
<td>Tensile strength</td>
<td>40 MPa</td>
<td>60 MPa</td>
</tr>
<tr>
<td>Density</td>
<td>0.92 g/cm³</td>
<td>0.94 g/cm³</td>
</tr>
<tr>
<td>Tensile strength</td>
<td>40 MPa</td>
<td>60 MPa</td>
</tr>
</tbody>
</table>

Thermosetting and thermoplastic polymers

The molecules within polymer chains are joined together by covalent bonds. In thermosetting polymers, these polymer chains are not joined together. In thermoplastic polymers, the polymer chains are joined to each other by covalent bonds (often called cross-links) (Table 10.15).

Table 10.15. Structures of thermosetting and thermoplastic polymers.

<table>
<thead>
<tr>
<th>Property</th>
<th>Formula</th>
<th>Polymer Chain</th>
<th>Uses of polymer</th>
</tr>
</thead>
<tbody>
<tr>
<td>Epoxy</td>
<td>HO(CH₂)₃OH</td>
<td>Polyetheramide</td>
<td>Coating, adhesives</td>
</tr>
<tr>
<td>Acrylic</td>
<td>CH₂=CH-CO₂H</td>
<td>Acrylonitrile</td>
<td>Polymers, adhesives, coatings</td>
</tr>
<tr>
<td>Polyethylene</td>
<td>(CH₂)₁₀</td>
<td>Polyethylene</td>
<td>Packaging, films, fibers</td>
</tr>
<tr>
<td>Polypropylene</td>
<td>(CH₂)₉-CH₂</td>
<td>Polypropylene</td>
<td>Plastics, fibers, films</td>
</tr>
</tbody>
</table>

Glass

Glass is used for windows and containers such as bottles. It is a very interesting material in that it is hard, see-through and unreactive. There are many different types of glass with slightly different properties and uses. Two of these are shown in Table 10.16.

Table 10.16. Properties of some glass.

<table>
<thead>
<tr>
<th>Property</th>
<th>Boro-silicate glass</th>
<th>soda-lime glass</th>
<th>kaolin glass</th>
<th>lead glass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Density</td>
<td>2.45 g/cm³</td>
<td>2.33 g/cm³</td>
<td>2.75 g/cm³</td>
<td>10.5 g/cm³</td>
</tr>
<tr>
<td>Tensile strength</td>
<td>40 MPa</td>
<td>60 MPa</td>
<td>60 MPa</td>
<td>50 MPa</td>
</tr>
<tr>
<td>Density</td>
<td>2.45 g/cm³</td>
<td>2.33 g/cm³</td>
<td>2.75 g/cm³</td>
<td>10.5 g/cm³</td>
</tr>
<tr>
<td>Tensile strength</td>
<td>40 MPa</td>
<td>60 MPa</td>
<td>60 MPa</td>
<td>50 MPa</td>
</tr>
</tbody>
</table>

Clay ceramics

Bricks and pottery such as plates are examples of clay ceramics. They are made by shaping wet clay that is then baked in a furnace at 980°C (Table 10.17). Clay ceramics are very hard, unreactive and resistant to heat.
Composites

Most composite materials are made from two or more different materials. The composite material has different properties from the materials in it. In composite materials, fibres or fragments of one material (called the reinforcement) are surrounded by a binder/matrix material that binds these together. For example, in fibreglass, glass fibres are bound together in a polymer. This and some other composites are described in Table 14.17.

<table>
<thead>
<tr>
<th>Composite material</th>
<th>Matrix</th>
<th>Reinforcement</th>
<th>Fibre</th>
<th>Properties</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fibreglass</td>
<td>Epoxy</td>
<td>Glass fibres</td>
<td>Highly resistant to fire, strong</td>
<td></td>
</tr>
<tr>
<td>Ceramic</td>
<td>Ceramic</td>
<td>Ceramic particles</td>
<td>Heat resistant, strong</td>
<td></td>
</tr>
<tr>
<td>Graphite</td>
<td>Graphite</td>
<td>Graphite flakes</td>
<td>Heat resistant, strong</td>
<td></td>
</tr>
<tr>
<td>Carbon fibre</td>
<td>Polymer</td>
<td>Carbon fibres or carbon nanotubes</td>
<td>Strong, lightweight</td>
<td></td>
</tr>
</tbody>
</table>

Test yourself

21. a) What is a polymer?
   b) Give two factors that can be changed to produce polymers with different properties.

22. a) What happens when a thermosetting polymer is heated?
   b) What happens when a thermoplastic polymer is heated?
   c) Explain why thermosetting and thermoplastic polymers set differently when moulded.

23. a) Which is an example of a silicone?
   b) Which is an example of a synthetic rubber?
   c) Which is an example of a synthetic rubber?
   d) Which is an example of a synthetic rubber?

Making fertilisers

Fertilisers are used to help crops grow. They supply nutrients that are essential for the growth of plants. The use of fertilisers increases the yield of crops significantly.

The world’s population reached 3 billion people in 1960. In 2011, it reached 7 billion people and it is forecast to keep rising in the future. We need to be able to provide enough food to feed everyone on the planet and many scientists argue that we could not do this without the use of fertilisers.

- The Haber process

Most fertilisers contain nitrogen compounds which are made from ammonia. Ammonia has the formula $NH_3$ and is made by the Haber process (Figure 10.18).

In the Haber process, nitrogen reacts with hydrogen. The nitrogen is obtained from the air. The hydrogen is made by reaction of steam with methane from natural gas or other sources (Figure 10.19).

Show you can...

Select a suitable material from the list below and give two reasons in each case why it is a suitable material.

- Thermosetting polymer
- Thermoplastic polymer
- Soluble glass
- Borax
- Carbon fibre composite

a) bonding glue
b) strong glue

c) using of heat insulators
The reaction between nitrogen and hydrogen reaches a state of dynamic equilibrium. The forward reaction is exothermic. In the equation, there are four reactant gas molecules but only two product gas molecules (Figure 10.39).

![Figure 10.39](image)

In addition, by recycling the leftover nitrogen and hydrogen, less raw materials are needed which saves costs.

### Production and uses of NPK fertilisers

Some water-soluble compounds of nitrogen (K), phosphorus (P) and potassium (K) are used as fertilisers to increase agricultural productivity. This means that they help crops grow and so increase the yield of the crops. Fertilisers that contain compounds of each of these three elements are called NPK fertilisers.

NPK fertilisers are a mixture of compounds in the correct proportions to give the desired N : P : K ratios. There are many different formulations of NPK fertilisers which contain different percentages of nitrogen, phosphorus and potassium. These different formulations are used for different crops and for different soil conditions (Figure 10.31).

Some of the compounds used in NPK fertilisers are shown in Table 10.39. They are all salts. Some of these salts are mixed on the ground, while others are made from reactions between acids and bases.

![Table 10.39](image)

Phosphate rock is often used as a raw material. It cannot be used directly as a fertiliser because it is insoluble in water. However, suitable salts can be made when the rock reacts with acids.

Many of the processes to make the compounds in NPK fertilisers are integrated. For example:

- Ammonium nitrate is made from the reaction of ammonia with nitric acid, but nitric acid itself is also made from ammonia (Figure 10.22).
Test yourself

1. Ammonia is made from reaction of nitrogen with hydrogen in the Haber process. Explain why this is important.
2. Write a balanced equation for the reaction.
3. Write the equilibrium constant expression for this reaction.
4. What is the rate of reaction? Explain this.
5. Why is it important? Explain.
6. From what raw materials is the nitrogen obtained? Explain.
7. How do ammoniums salts form from the reaction of ammonia and hydrogen in the Haber process?
8. When does the reaction become irreversible?

Chapter review questions

1. Ammonia is made from reaction of nitrogen with hydrogen in the Haber process.
2. Write a balanced equation for the reaction.
3. Write the equilibrium constant expression for this reaction.
4. What is the rate of reaction? Explain this.
5. Why is it important? Explain.
6. From what raw materials is the nitrogen obtained? Explain.
7. How do ammonium salts form from the reaction of ammonia and hydrogen in the Haber process?
8. When does the reaction become irreversible?
Practice questions

1. Which one of the following is NOT true of the conditions under which the Haber process is conducted? [1 mark]
   a. The temperature is approximately 400°C.
   b. The raw materials are at a high pressure.
   c. The reactants are mixed in the ratio of 1 part nitrogen to 1 part hydrogen.
   d. Ammonia is a product of the reaction.
   e. The process is carried out in a autoclave.

2. Which of the following could be used to obtain a sample of pure water from sea water? [1 mark]
   a. chromatography
   b. crystallization
   c. distillation
   d. filtration
   e. precipitation

3. Nitrogenous fertilizers contain organic compounds such as ammonia which is produced when ammonia reacts with nitric acid.

4. In the United Kingdom, over 30 million mobile phone users discard over 2.5 million handsets every year. Only a small percentage of recycled phones are recycled for raw materials, turning the waste phones into materials for use in new devices. How many kilograms of waste mobile phones are collected in a year? [1 mark]

5. State the effect of increasing temperature on the yield of ammonia at constant pressure. [1 mark]

6. State the effect of increasing pressure on the yield of ammonia at constant temperature. [1 mark]

7. Choose the correct formula for copper(II) carbonate. [1 mark]
   a. CuCO₃
   b. Cu₂CO₃
   c. Cu₂(CO₃)₂
   d. CuCO₃·H₂O
   e. Cu₂(CO₃)₃·3H₂O

8. The graph below shows the percentage yield of ammonia changes with pressure and temperature. Use the graph to answer the following questions.

   a. State the trend of the percentage yield of ammonia with pressure. [1 mark]
   b. State the trend of the percentage yield of ammonia with temperature. [1 mark]
   c. What is the effect of pressure on the yield of ammonia? [1 mark]
   d. State the pressure at which the yield of ammonia is maximum. [1 mark]

9. State the advantage and one disadvantage of recuperation as a method of disposal of plastics. [1 mark]

10. What is meant by the term alloy? [1 mark]

11. What is the ultimate tensile strength of steel? [1 mark]

12. Why is the percentage of nickel in stainless steel typically 10-20%? [1 mark]
1. Copper can be found in the Earth’s crust as native copper. Describe where copper is most likely to be found in the Earth’s crust.
2. Copper oxide is a common mineral. Describe the physical and chemical properties of copper oxide.
3. Copper is used in a variety of applications. List three common uses of copper and explain why it is chosen for these applications.
4. Copper is an element with distinct characteristics. Describe the electron configuration of copper and explain how it affects its physical properties.
5. Copper is a good conductor of electricity. Explain why copper is chosen for electrical wiring and explain the role of copper oxide in electrical conductivity.
6. Copper is a hard material. Explain why copper is used for making pipes and plumbing fixtures.
7. Copper is a reactive metal. Explain how copper reacts with other substances and give examples of specific reactions that copper undergoes.
8. Copper is a common metal. Describe the common physical and chemical properties of copper and explain why it is chosen for these applications.
9. Copper is a valuable metal. Explain why copper is used for making jewelry and explain the role of copper oxide in jewelry making.
10. Copper is a good conductor of heat. Explain why copper is chosen for cooking utensils and explain the role of copper oxide in heat conductivity.
Questions

1. Two students carried out experiments to find the percentage by mass of nitrogen in ammonia. The first student got 23%, the second 25%. Which of the students was more accurate?
2. Calculate the mass of each student’s sample, and explain your answer.
3. Comment on the reproducibility of the result for each student.
4. Why do the results differ when each student repeats them?
5. Is the other of the students have a systematic error or an experimental error? Explain your answer.

A: Scale measured the mass of ammonium nitrate (NH₄NO₃) in a 250 g sample of this substance. The sample was placed on a balance, and the mass was recorded as 250.0 g. What is the accuracy of the measurement?
B: A mixture of ammonium nitrate and sugar was placed on a balance, and the mass was recorded as 300.0 g. What is the accuracy of the measurement?

C: A mixture of ammonium nitrate and sugar was placed on a balance, and the mass was recorded as 250.0 g. What is the reproducibility of the measurement?

D: A mixture of ammonium nitrate and sugar was placed on a balance, and the mass was recorded as 250.0 g. What is the reproducibility of the measurement?

E: A mixture of ammonium nitrate and sugar was placed on a balance, and the mass was recorded as 250.0 g. What is the reproducibility of the measurement?

2. In an experiment to determine the density of a liquid, a 25 ml graduated cylinder was used. The cylinder was filled with water, and the mass was recorded as 25.0 g. What is the accuracy of the measurement?

3. In an experiment to determine the density of a liquid, a 25 ml graduated cylinder was used. The cylinder was filled with water, and the mass was recorded as 25.0 g. What is the reproducibility of the measurement?

4. In an experiment to determine the density of a liquid, a 25 ml graduated cylinder was used. The cylinder was filled with water, and the mass was recorded as 25.0 g. What is the reproducibility of the measurement?

5. In an experiment to determine the density of a liquid, a 25 ml graduated cylinder was used. The cylinder was filled with water, and the mass was recorded as 25.0 g. What is the reproducibility of the measurement?

This chapter is called Formulae and Equations as it is a quick way of writing and understanding chemical formulae and equations. What are some of the challenges you face when writing and using chemical formulae and equations?
Writing formulae

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 11.1).
  - 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).

- **Compounds**
  - It is very useful to know the formula of some common compounds. Some are listed in Table 11.2.

### Table 11.1

<table>
<thead>
<tr>
<th>Compound</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>H₂O</td>
<td>Water</td>
</tr>
<tr>
<td>CO₂</td>
<td>Carbon dioxide</td>
</tr>
<tr>
<td>NaCl</td>
<td>Sodium chloride</td>
</tr>
<tr>
<td>NH₄Cl</td>
<td>Ammonium chloride</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 11.3 and 11.4.

#### Table 11.3: Positive ions

<table>
<thead>
<tr>
<th>Ion</th>
<th>Charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li⁺</td>
<td>1</td>
</tr>
<tr>
<td>Na⁺</td>
<td>1</td>
</tr>
<tr>
<td>K⁺</td>
<td>1</td>
</tr>
<tr>
<td>Ca²⁺</td>
<td>2</td>
</tr>
<tr>
<td>Mg²⁺</td>
<td>2</td>
</tr>
</tbody>
</table>

#### Table 11.4: Negative ions

<table>
<thead>
<tr>
<th>Ion</th>
<th>Charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl⁻</td>
<td>1</td>
</tr>
<tr>
<td>O²⁻</td>
<td>2</td>
</tr>
<tr>
<td>SO₄²⁻</td>
<td>2</td>
</tr>
<tr>
<td>NO₃⁻</td>
<td>1</td>
</tr>
<tr>
<td>PO₄³⁻</td>
<td>3</td>
</tr>
</tbody>
</table>

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 chloride (NaCl)

- Sodium has a charge of +1 (Na⁺)
- Chloride has a charge of -1 (Cl⁻)
- The formula is NaCl

#### Calcium carbonate (CaCO₃)

- Calcium has a charge of +2 (Ca²⁺)
- Carbonate has a charge of -2 (CO₃²⁻)
- The formula is CaCO₃

Some ions contain atoms of different elements. Examples include sulfate (SO₄²⁻), hydrogen (H⁺) 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.
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 11.6 gives some general guidelines on this.

- **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 11.7).

- **Acid-base character of oxides**
  Most metal oxides are basic (Table 11.8). For example, calcium oxide (CaO) is used as a base to neutralise acidic soil or farms. Most non-metal oxides are acidic. For example, carbon dioxide (CO₂) dissolves in rain water to make rainwater slightly acidic.

---

### Table 11.5 - Formula and equations

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>Charge</th>
<th>Charge</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>H₂O</td>
<td>2⁺</td>
<td>2⁻</td>
<td>H₂O⁺₂H⁻</td>
</tr>
<tr>
<td>Magnesium chloride</td>
<td>MgCl₂</td>
<td>2⁺</td>
<td>2⁻</td>
<td>MgCl₂</td>
</tr>
<tr>
<td>Magnesium sulfate</td>
<td>MgSO₄</td>
<td>2⁺</td>
<td>4⁻</td>
<td>MgSO₄</td>
</tr>
<tr>
<td>Copper sulfate</td>
<td>CuSO₄</td>
<td>2⁺</td>
<td>4⁻</td>
<td>CuSO₄</td>
</tr>
<tr>
<td>Ammonium sulfate</td>
<td>(NH₄)₂SO₄</td>
<td>2⁺</td>
<td>4⁻</td>
<td>(NH₄)₂SO₄</td>
</tr>
<tr>
<td>Calcium carbonate</td>
<td>CaCO₃</td>
<td>2⁺</td>
<td>3⁻</td>
<td>CaCO₃</td>
</tr>
<tr>
<td>Aluminum oxide</td>
<td>Al₂O₃</td>
<td>2⁺</td>
<td>3⁻</td>
<td>Al₂O₃</td>
</tr>
<tr>
<td>Bismuth trioxide</td>
<td>Bi₂O₃</td>
<td>3⁺</td>
<td>3⁻</td>
<td>Bi₂O₃</td>
</tr>
</tbody>
</table>

### Table 11.6 - Structure type

<table>
<thead>
<tr>
<th>Structure type</th>
<th>Description of structure</th>
<th>Molecule or compound base this structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ionic</td>
<td>Made up of charged particles (e.g., Na⁺, Cl⁻)</td>
<td>NaCl (sodium chloride)</td>
</tr>
<tr>
<td>Molecular</td>
<td>Made up of individual molecules</td>
<td>CO₂ (carbon dioxide)</td>
</tr>
<tr>
<td>Giant covalent</td>
<td>Lattice of atoms joined by covalent bonds</td>
<td>Diamond (C)</td>
</tr>
<tr>
<td>Non-polar</td>
<td>Lattice of electrically neutral ions</td>
<td>NaCl (sodium chloride)</td>
</tr>
<tr>
<td>Polar</td>
<td>Lattice of electrically neutral ions with a charge difference</td>
<td>NaCl (sodium chloride)</td>
</tr>
</tbody>
</table>

### Table 11.7 - Acids, bases, alkalis and salts

<table>
<thead>
<tr>
<th>Acid</th>
<th>Formula</th>
<th>Base</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td>Cl⁻</td>
<td>NaOH</td>
<td>Na⁺ OH⁻</td>
</tr>
<tr>
<td>H₂SO₄</td>
<td>SO₄²⁻</td>
<td>Ca(OH)₂</td>
<td>Ca²⁺ (OH)₂⁻</td>
</tr>
<tr>
<td>H₂CO₃</td>
<td>CO₃²⁻</td>
<td>Na₂CO₃</td>
<td>Na⁺₂ CO₃⁻</td>
</tr>
</tbody>
</table>

---

### Test yourself

1. Write the formula of each of the following elements and compounds:
   - Oxygen (O₂)
   - Carbon (C) and Carbon dioxide (CO₂)
   - Sodium (Na) and Sodium chloride (NaCl)

2. Write the formula of each of the following ionic compounds:
   - Magnesium nitrate (MgNO₃)
   - Potassium carbonate (K₂CO₃)
   - Alkaline compounds (e.g., KOH, NaOH)

---

Example

Mg(OH)₂ contains magnesium ions (Mg²⁺) and hydroxide ions (OH⁻). There must be the same number of positive and negative charges, one magnesium ion has a charge of two positive charges for every two OH⁻ (two times two negative charges).
### Common Reactions

Here are some common reactions that are useful to know. Many of these reactions are covered in more detail in later chapters. Some reactions differ with ionic strength, and proteins often exhibit ionic effects.

#### Table 11.1 Formulas and Equations

- **Table 11.1a** Formulas of Common Reagents

<table>
<thead>
<tr>
<th>Reagent</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium</td>
<td>Na⁺</td>
</tr>
<tr>
<td>Potassium</td>
<td>K⁺</td>
</tr>
<tr>
<td>Calcium</td>
<td>Ca²⁺</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Mg²⁺</td>
</tr>
<tr>
<td>Chloride</td>
<td>Cl⁻</td>
</tr>
<tr>
<td>Nitrate</td>
<td>NO₃⁻</td>
</tr>
<tr>
<td>Sulfate</td>
<td>SO₄²⁻</td>
</tr>
<tr>
<td>Phosphate</td>
<td>PO₄³⁻</td>
</tr>
</tbody>
</table>

- **Table 11.1b** Equations for Reactions

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Neutralization</td>
<td>Na⁺ + OH⁻ → NaOH</td>
</tr>
<tr>
<td>Precipitation</td>
<td>Ag⁺ + Cl⁻ → AgCl</td>
</tr>
<tr>
<td>Oxidation</td>
<td>Fe²⁺ + H₂O₂ → Fe³⁺ + H₂O</td>
</tr>
<tr>
<td>Reduction</td>
<td>Cu²⁺ + Fe → Cu + Fe²⁺</td>
</tr>
</tbody>
</table>

- **Table 11.1c** Other Chemical Reactions

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Displacement</td>
<td>Na⁺ + Cl⁻ → NaCl</td>
</tr>
<tr>
<td>Explosion</td>
<td>2Na + O₂ → Na₂O₂</td>
</tr>
<tr>
<td>Combustion</td>
<td>C + O₂ → CO₂</td>
</tr>
</tbody>
</table>

- **Table 11.1d** Sample Experiments and Calculations

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Calculation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Titration</td>
<td>V₁ₙ₁₀₀₀ × C₁ = V₂ₙ₁₀₀₀ × C₂</td>
</tr>
<tr>
<td>Enzyme Assay</td>
<td>Aₚ = A₀ – Aₑ</td>
</tr>
</tbody>
</table>

### Calculations

**Example Calculation:**

- **Preparation of a Standard Solution:**

  - **Volume of Solution:** V = 100 mL
  - **Concentration of Solution:** C = 0.1 M
  - **Result:** V × C = 10 mL

- **Calculation of Molarity:**

  - **Moles of Solute:** n = 0.5 mol
  - **Volume of Solution:** V = 250 mL
  - **Result:** n/V = 0.002 M

- **Titration Value:**

  - **Initial Volume:** V₁ = 20 mL
  - **Final Volume:** V₂ = 40 mL
  - **Result:** V₂ - V₁ = 20 mL

- **Enzyme Activity:**

  - **Activity:** A = 2.5 U/mL
  - **Time:** t = 15 min
  - **Result:** A × t = 37.5 U min/mL
Balancing equations

Most equations show the names of the reactants and products in a reaction. A balanced 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 11.13.

### Table 11.13

<table>
<thead>
<tr>
<th>Equation</th>
<th>Reaction</th>
<th>Balanced equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Oxygen + hydrogen → water</td>
<td>$2 \text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}$</td>
<td>$2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$</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: Write 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 $\text{H}_2\text{O}$ to $\text{H}_2\text{O}_2$ in Exercise 1 to balance the $\text{H}_2$ atoms because it is water that is formed and that has the formula $\text{H}_2\text{O}$ and not $\text{H}_2\text{O}_2$.

#### Step 5: Write out the final balanced equation.

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.
Test yourself

1. Magnesium reacts with sulfuric acid as shown:
   \[ \text{Mg + H}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{H}_2 \]
   a) What does the \( \text{Mg} \) mean?
   b) What does the \( \text{H}_2 \) mean?
   c) What does the \( \text{SO}_4 \) mean?

2. Balance the following equations:
   a) \( \text{K}_2\text{SO}_4 \rightarrow \text{K}_2\text{O} + \text{SO}_3 \)
   b) \( \text{CaCl}_2 \rightarrow \text{CaCl}_2 + \text{H}_2 \)
   c) \( \text{CaF}_2 \rightarrow \text{CaF}_2 + \text{H}_2 \)
   d) \( \text{Na}_2\text{O}_3 \rightarrow \text{Na}_2\text{O}_3 + \text{H}_2 \)
   e) \( \text{MgCl}_2 + \text{H}_2 \rightarrow \text{MgCl}_2 \)

3. Write a balanced equation for each of the following reactions:
   a) Sodium + oxygen \rightarrow sodium oxide
   b) Potassium + oxygen \rightarrow potassium oxide
   c) Calcium + oxygen \rightarrow calcium oxide
   d) Sodium + water \rightarrow sodium hydroxide + hydrogen
   e) Magnesium + water \rightarrow magnesium oxide + hydrogen

In a solid ionic compound, the positive and negative ions are bound to each other strongly in a lattice. When it dissolves in water, the ions separate and become surrounded by water molecules (Figure 11.2).

When ionic compounds dissolve in water, they do not break down to become atoms. Instead, they dissociate into their constituent ions (Figure 11.1). The ions are carried in solution as 'ions'

To balance the equation for the reaction between an acid and a base, we write an equation for the reaction and balance the equation for the reaction.

The reaction of acids with alkalis:

When sulfuric acid reacts with sodium hydroxide solution, it is not only the hydrogen ions from the sulfuric acid and the hydroxide ions from the sodium hydroxide that react. These hydroxide ions and hydronium ions react to form water (Figure 11.2). 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.

Table 11.10

<table>
<thead>
<tr>
<th>Reaction</th>
<th>What reacts</th>
<th>ions that do not react</th>
<th>Ionic equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium hydroxide (aq) + \text{H}_2\text{SO}_4 (aq)</td>
<td>\text{Sodium from } \text{NaOH}</td>
<td>\text{H}^+ from \text{H}_2\text{SO}_4</td>
<td>\text{Sodium from } \text{NaOH}  \rightarrow \text{Na}^+ + \text{OH}^- + \text{H}_2\text{SO}_4 \rightarrow \text{H}^+ + \text{SO}_4^{2-}</td>
</tr>
<tr>
<td>\text{Potassium hydroxide} (aq) + \text{H}_2\text{SO}_4 (aq)</td>
<td>\text{Potassium from } \text{KOH}</td>
<td>\text{H}^+ from \text{H}_2\text{SO}_4</td>
<td>\text{Potassium from } \text{KOH} \rightarrow \text{K}^+ + \text{OH}^- + \text{H}_2\text{SO}_4 \rightarrow \text{H}^+ + \text{SO}_4^{2-}</td>
</tr>
<tr>
<td>\text{Calcium hydroxide} (aq) + \text{H}_2\text{SO}_4 (aq)</td>
<td>\text{Calcium from } \text{Ca(OH)}_2</td>
<td>\text{H}^+ from \text{H}_2\text{SO}_4</td>
<td>\text{Calcium from } \text{Ca(OH)}_2 \rightarrow \text{Ca}^2+ + 2\text{OH}^- + \text{H}_2\text{SO}_4 \rightarrow \text{H}^+ + \text{SO}_4^{2-}</td>
</tr>
<tr>
<td>\text{Magnesium hydroxide} (aq) + \text{H}_2\text{SO}_4 (aq)</td>
<td>\text{Magnesium from } \text{Mg(OH)}_2</td>
<td>\text{H}^+ from \text{H}_2\text{SO}_4</td>
<td>\text{Magnesium from } \text{Mg(OH)}_2 \rightarrow \text{Mg}^2+ + 2\text{OH}^- + \text{H}_2\text{SO}_4 \rightarrow \text{H}^+ + \text{SO}_4^{2-}</td>
</tr>
</tbody>
</table>


Precipitation reactions

Precipitate: A solid formed when two solutions of ionic compounds are mixed. For example, when barium chloride solution is mixed with sodium sulfate solution, the barium ions from one solution combine with the sulfate ions from the other solution to form barium sulfate which is insoluble in water (Figure 11.3). The chloride ions and the sodium ions do not react and are not included in the ionic equation as they are spectator ions.

\[ \text{Ba}^{2+} (aq) + \text{SO}_4^{2-} (aq) \rightarrow \text{BaSO}_4 (s) \]

Figure 11.3

The ionic equation for this reaction can be written as:

\[ \text{Ba}^{2+} (aq) + \text{SO}_4^{2-} (aq) \rightarrow \text{BaSO}_4 (s) \]

There are many examples of precipitation reactions, including reactions that are used to test for ions in solution. When writing the ionic equation, it is very important to get the formula of the precipitate correct. This also helps you to balance the equation. In the next example, the formula of copper hydroxide is Cu(OH)_2, and so one Cu^{2+} ion and two OH^- ions are needed to balance the equation.

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 ion in the copper sulfate (Figure 11.4). The sulfate ions are spectator ions and so do not appear in the ionic equation.

\[ \text{Zn}(s) + \text{Cu}^{2+} (aq) \rightarrow \text{Zn}^{2+} (aq) + \text{Cu}(s) \]

Figure 11.4

The ionic equation for this reaction can be written as:

\[ \text{Zn}(s) + \text{Cu}^{2+} (aq) \rightarrow \text{Zn}^{2+} (aq) + \text{Cu}(s) \]

This, and some other examples of displacement reactions are shown in Table 11.12. In these reactions it helps to use the electric charges of the ions to balance the equations. For example, in the second example, in the table, two \text{Ag}^+ ions are needed gaining an overall +2 charge on the left side of the equation to balance with the +2 charge on the \text{Fe}^{2+} ion on the right side of the equation.
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., $A^+ + e^- \rightarrow A$)
- negative ions lose electrons (e.g., $S^2- - 2e^- \rightarrow S$)

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 $S^2-$ ions to form $S$ can be written as:

$$S^2- - 2e^- \rightarrow S$$

Some elements contain diatomic molecules (e.g., $H_2$, $O_2$, $Cl_2$, $Br_2$, $I_2$). When balancing half-equations that produce these elements, two ions are needed to make one diatomic molecule.

For example, when $H_2$ is formed from $H^+$ ions, two $H^+$ ions are needed which both gain one electron and so two electrons are gained altogether:

$$2H^+ + 2e^- \rightarrow H_2$$

When $O_2$ is formed from $O^-$ ions, two $O^-$ ions are needed which both lose two electrons and so four electrons are lost altogether:

$$2O^- - 4e^- \rightarrow O_2$$

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):

$$Al^{3+} + 3e^- \rightarrow Al$$

Reduction takes place when a substance loses electrons. Reduction takes place when a substance loses electrons. One way to remember this is the phrase O3 R3S (Figure 11.15).

Electrolysis

Half-equations can be written for the reactions that take place at each electrode in electrolysis. Some examples are shown in Table 11.14.
**Glossary**

Acerator: Reactor used to produce high-voltage electric power.

Addition reaction: A reaction in which one or more of the atoms of a reagent is added to one or more of the atoms of another reagent to form a new compound.

Addition polymerization: A type of polymerization in which monomers are added to each other to form a polymer.

Alkyl: A group derived from an alkane, consisting of one or more carbon atoms and hydrogen atoms.

AlkB: A bacterial enzyme that repairs DNA by removing damaged bases.

Alkene: An unsaturated hydrocarbon that contains at least one carbon-carbon double bond.

Alcohol: A compound that contains a hydroxyl group (-OH) and is soluble in water.

Alkaline: A solution that is basic and has a pH greater than 7.

Alkaline earth metals: A group of elements with atomic numbers 12 to 18, including beryllium, magnesium, calcium, strontium, barium, and radium.

Alloy: A metal that is a combination of two or more metals, typically with a metallic bond.

Aluminum: A soft, lightweight, flexible metal that is highly reflective and has a low electrical and thermal conductivity.

**Table 1.14**

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Product</th>
<th>Reaction</th>
<th>Product</th>
</tr>
</thead>
<tbody>
<tr>
<td>2Na + 2Cl₂ → 2NaCl</td>
<td>Sodium chloride</td>
<td>2H₂ + O₂ → 2H₂O</td>
<td>Water</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(II) sulfate solution, the overall ionic equation is:

\[ Zn^{2+} + 2e^- \rightarrow Zn \quad \text{and} \quad Cu^{2+} + 2e^- \rightarrow Cu \]

The two half equations for this are:

\[ Zn + CuSO₄ → ZnSO₄ + Cu \]

**Test yourself**

1. Write a balanced half equation for each of the following reactions:
   - a. \( \text{Mg}^{2+} + \text{H}_2 \rightarrow \text{Mg} + \text{H}_2\text{O} \)
   - b. \( \text{Fe} + \text{CuSO}_4 \rightarrow \text{FeSO}_4 + \text{Cu} \)
   - c. \( \text{H}_2 + \text{Cl}_2 \rightarrow \text{H}_2\text{Cl}_2 \)
   - d. \( \text{Al} + \text{FeCl}_3 \rightarrow \text{AlCl}_3 + \text{Fe} \)
   - e. \( \text{Cu} + \text{HCl} \rightarrow \text{CuCl}_2 + \text{H}_2 \)

2. Write a balanced half equation for each of the following reactions:
   - a. \( \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \)
   - b. \( \text{Mg}^{2+} + 2e^- \rightarrow \text{Mg} \)
   - c. \( \text{Al}^{3+} + 3e^- \rightarrow \text{Al} \)

3. Write a balanced half equation for each of the following reactions:
   - a. \( \text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu} \)
   - b. \( \text{Cu} + \text{HCl} \rightarrow \text{CuCl}_2 + \text{H}_2 \)
   - c. \( \text{Al} + \text{H}_2\text{O} \rightarrow \text{Al}_2\text{O}_3 + \text{H}_2 \)

4. Write a balanced half equation for each of the following reactions:
   - a. \( \text{Zn}^{2+} + 2e^- \rightarrow \text{Zn} \)
   - b. \( \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \)
   - c. \( \text{Al}^{3+} + 3e^- \rightarrow \text{Al} \)