

Richard Harwood and Ian Lodge
Cambridge IGCSE®
Chemistry
Coursebook

Fourth edition



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Introduction

Chemistry is a laboratory science: its subject material and theories are based on experimental observation. However, its scope reaches out beyond the laboratory into every aspect of our lives – to our understanding of the nature of our planet, the environment we live in, the resources available to us and the factors that affect our health.

This book thoroughly covers the Cambridge International Examinations (CIE) Chemistry syllabus and includes features which are aimed at helping you grasp the concepts and detail involved. The areas that cover the Core and Supplement material of the syllabus are clearly marked (the Supplement material having a purple bar like the one here in the margin) so that you can see which topics will be tested on each exam paper that you will take. The topic summaries, questions and end-of-chapter questions are also clearly marked so that you can pick out, study and revise the material relevant to the ‘core’ and ‘extended’ papers.

The first chapter of the book serves to set chemistry in its broader context and as such contains material that ‘sets the scene’ as well as syllabus material. At various points in this and other chapters there is material that provides and develops some of the context in which chemical ideas are important. These are areas such as:

- ◆ the importance of chemistry to life, and the nature of the universe (Chapter 1)
- ◆ renewable and non-renewable resources (Chapter 1)
- ◆ our need to develop alternative energy sources (Chapter 11).

The introduction to each chapter aims to highlight some of the more novel aspects of chemistry – from unusual alloys and the visualisation of the bonding in molecules to the analytical laboratory on the surface of Mars.

Features of the book and the Student CD-ROM

The book is divided into broad **chapters** covering important areas of the syllabus. These chapters are then divided into different sections to help you manage your

understanding of the ideas involved. At the end of each section there are short **questions** to help you check that you have followed the ideas covered. The **answers** to these short questions are provided at the end of the book to help you with this. Included in the text are a series of **study tips** and **key definitions**. These highlight important areas of learning and useful approaches to a particular topic.

Each chapter finishes with a **summary** of Core and Supplement material to help you particularly with your revision. This is followed by a selection of **end-of-chapter questions** which are there to help you become familiar with the style of question set in each examination.

Answering questions is a great way to get to grips with each of the topics. However, it is not the only way! The **Student CD-ROM** provides information on revision skills and resources available on the internet to help with your study of chemistry. A copy of the syllabus is provided on the CD-ROM, which shows where the different topics are covered in the book – and you can use this interactively as a checklist during revision. ‘Mind-mapping’ ideas and other revision strategies are discussed on the CD-ROM, and we hope that you can find ideas that will help you study in the most personally effective way.

An important feature which appears both in the book and on the Student CD-ROM is the **glossary**. The terms included in the glossary are highlighted in the text in **dark red bold**. Do use this resource in addition to the text in order to help you understand the meaning of chemical terms. But more than that, it is important that you can express your ideas clearly in an exam – that is why we have included so many practice questions in the book and in the **practice tests** that appear on the CD-ROM. It is also why we have tried to cover the ideas in each chapter thoroughly in our wording. The information boxes and ‘key definitions’ placed throughout the chapters are there to help you learn how to summarise your knowledge in an effective and clear way. Chemistry, and science in general, can often use certain words in a very precise way, so it is important to read carefully and get used to writing down your answers clearly.

Practical work

We began by saying that chemistry was a practical science and, in this edition, we have included **Activities** throughout the chapters, which we hope will encourage your enjoyment of the practical aspect of the subject. Worksheets for these practicals are included on the Student CD-ROM. In addition, we have aimed to help your preparation for the practical element of the exam in various ways:

- ◆ Chapter 12 of the book gives a summary of the different ways that practical work is assessed and some exemplar questions.
- ◆ There are practice ‘alternative to practical’ papers (Paper 6) on the Student CD-ROM.
- ◆ The separate **Student Workbook** contains exercises involving practice at the key skills of writing up your

observations and making deductions from your results. Included there are methods that you can use to assess (by yourself, and with your teacher) how well you are developing your data handling and presentation skills.

Chemistry is an important, exciting and challenging subject that impacts on every aspect of our lives. As we face the challenges of the future, the chemical ‘angle’ on things will figure in our thinking, whatever future course we personally take in our careers. We hope that this book will help you enjoy chemistry, give you some understanding of the ideas involved and help you be successful in the IGCSE course.

Richard Harwood

Ian Lodge

1 Planet Earth

In this chapter, you will find out about:

- ◆ the water cycle
- S** ◆ the carbon cycle
- ◆ the nitrogen cycle
- ◆ the composition and uses of the gases in the air
- S** ◆ the separation of air into its components
- ◆ the sources of air pollution
- ◆ the problems of air pollution, and their solution
- ◆ ‘greenhouse gases’ and climate change
- ◆ water treatment and sewage treatment
- ◆ the pollution of water
- ◆ metal ores and limestone
- ◆ fossil fuels and the problems they cause
- ◆ alternative sources of energy
- ◆ hydrogen as a fuel
- S** ◆ the hydrogen fuel cell.

A brief history of the Earth



Figure 1.1 A satellite image over Africa: one view of the ‘blue marble’.

The Earth is a ball of rock orbiting a star along with a group of other planets (Figure 1.1). The star is one of many billions of stars in a galaxy which, in turn, is one of many billion galaxies in a constantly expanding Universe. As such, the Earth is

unremarkable. It is the chemicals which make up the Earth and the ways in which they interact with each other that make life on Earth possible.

At the start, the Earth was a ball of molten rock. The surface solidified to a solid crust as it cooled and contracted, and cracks appeared. Volcanoes shot molten rock and gases from this surface and the first atmosphere (mainly carbon dioxide and water vapour) was formed.

Condensing water vapour fell back to the surface and, over many millions of years, plant life developed in these warm, shallow seas. The plants used carbon dioxide in **photosynthesis** and, crucially, put oxygen into the atmosphere. Once sufficient oxygen was present, animal life began to evolve. Nitrogen entered the atmosphere from bacteria. Because nitrogen is an unreactive gas, it was not removed and it has built up to a large percentage of the atmosphere.

The development of plant and animal life over many millions of years has led to the Earth’s present balance of chemicals. The activity of humans is now altering this chemical balance and we are rapidly using up many of the Earth’s natural resources.

Study tip

This chapter provides a context for the chemistry that you study. As such, it makes some general comments about the origins of the Earth and the nature of the natural resource cycles that occur. The list at the start of this chapter is similar to those given at the beginning of each chapter. It gives you an idea of the material in the chapter that is contained in the syllabus and that can therefore be examined.

- ◆ As it cools, it changes back into liquid water and forms clouds (tiny water droplets).
- ◆ As the water droplets stick together, rain clouds are formed and the water falls back to the surface as rain, snow or hail.
- ◆ Water then either flows back to the sea or is taken in by plants, which put it back into the atmosphere through their leaves.
- ◆ We use the water by trapping it on its way back to the sea.

The carbon cycle

Carbon is only the twelfth most common element in the Earth, making up less than 1% of the crust. It is, however, very important to us. Without carbon, life would not exist. The way in which carbon moves around in the **carbon cycle** is vital to all life (Figure 1.3).

The source of the carbon in the cycle is carbon dioxide in the atmosphere. Only about 0.04% of the atmosphere is carbon dioxide.

Carbon dioxide leaves the atmosphere in the following ways:

- ◆ Green plants take carbon dioxide and water, combining them together to form glucose and oxygen. This process uses energy from the Sun and is called **photosynthesis**. The word equation for the reaction is:

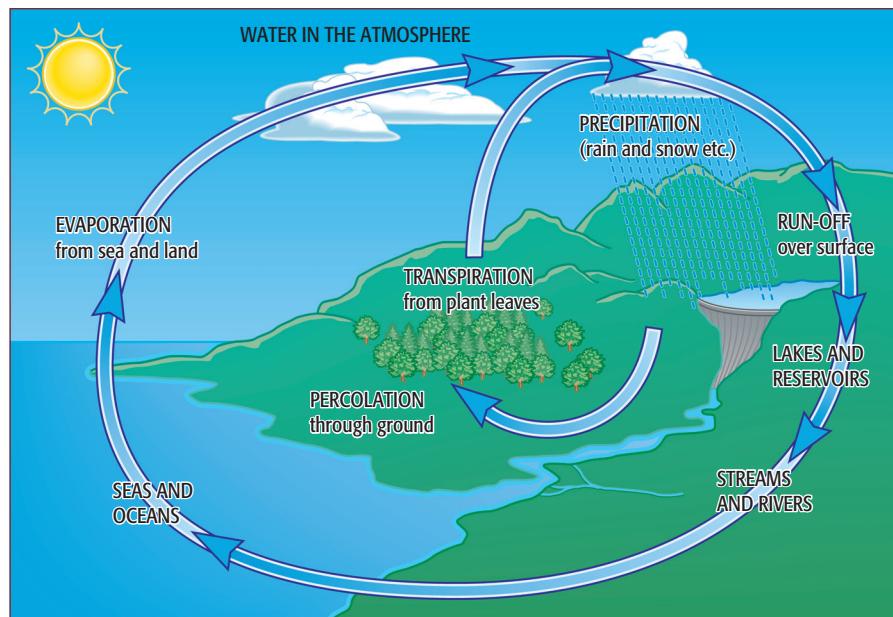


Figure 1.2 The water cycle.

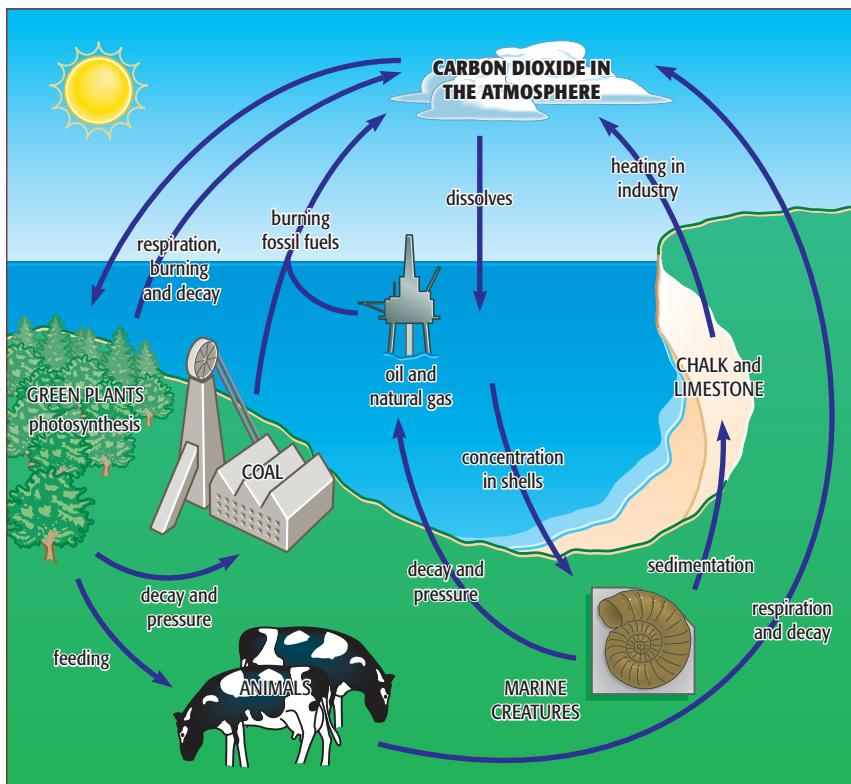


Figure 1.3 The carbon cycle.

- ◆ Carbon dioxide dissolves in water (mainly seawater), where it is used by animals and plants. Plants use it in photosynthesis; animals use it to make their shells.

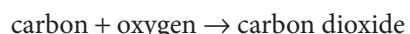
This is what happens to the carbon once it has been captured from the atmosphere:

- ◆ The plants are eaten by animals.
- ◆ Animals and plants die and rot away, or they are buried and slowly (over millions of years) are fossilised.
- ◆ Tiny sea creatures die and their bodies fall to the bottom of the sea where they slowly (over millions of years) change to limestone.

These are the ways in which carbon dioxide is put back into the atmosphere:

- ◆ Animals and plants 'breathe out' carbon dioxide when they respire. The process of **respiration** uses oxygen from the air and releases carbon dioxide:
$$\text{glucose} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}$$
- ◆ When plants and animals decay after death, carbon dioxide is produced.

- ◆ Wood can be burnt. This **combustion** produces carbon dioxide:



- ◆ Fossilised plants and animals form **fossil fuels** (coal, oil and gas); these produce carbon dioxide when they are burnt.
- ◆ Limestone produces carbon dioxide when it is heated in industry and when it moves back below the Earth's crust.

The problem we face is balancing the amount of carbon dioxide being added to the atmosphere with the amount being taken out by plants and the oceans (Figure 1.4, overleaf).

The nitrogen cycle

Nitrogen is essential for plant growth and therefore for the life of animals (Figure 1.5, overleaf). There is plenty of nitrogen in the atmosphere (78%) but it is unreactive and so it is difficult to get it into the soil for plants to use.

Plants generally get their nitrogen from nitrates in the soil and animals get theirs from eating plants.

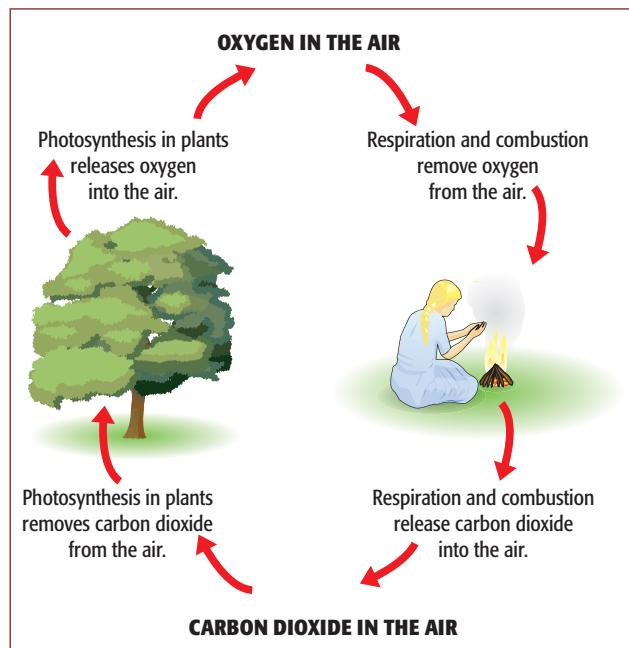


Figure 1.4 Maintaining the levels of oxygen and carbon dioxide in the air.

When plants and animals die and decay, bacteria help the decomposition and nitrogen is returned to the soil.

There are also bacteria that live in the roots of some plants (e.g. beans and clover) that can 'fix' nitrogen

from the atmosphere which the plants can then use. This process is called **nitrogen fixation**.

During thunderstorms, the very high temperature of the lightning provides enough energy to cause atmospheric nitrogen and oxygen to react with water in the atmosphere to form nitric acid. When this falls with rain, it forms nitrates in the soil. Nitrogen is also taken from the air by the chemical industry when fertiliser is made by the Haber process.

Taken together, these processes form the **nitrogen cycle** (Figure 1.5).

These three major cycles – of water, carbon and nitrogen – together with the rock cycle interlink and, between them, provide us with the resources we need.

The Earth's resources

In human terms, **resources** are materials we get from the environment to meet our needs. Some are the basic material resources we and other organisms need to keep alive; others are materials from which we obtain energy, or substances useful for our civilised way of life. Chemistry helps us to understand how the basic resources sustain our life. It also provides the methods of extraction and use of other resources.

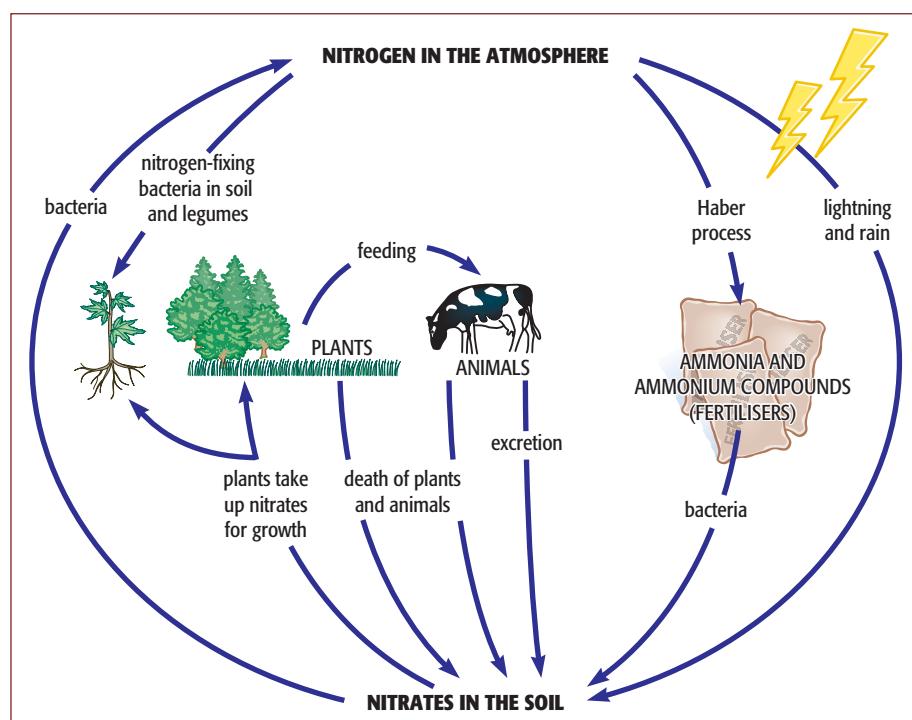


Figure 1.5 The nitrogen cycle.

Material resources can be broadly subdivided into **renewable**, **potentially renewable** and **non-renewable resources**, based on our short human timescale.

- ◆ **Non-renewable resources** are those that exist in a fixed quantity in the Earth's crust – for example, metallic and non-metallic minerals and fossil fuels. They were formed over millions of years and are being used up much faster than they are being formed.
- ◆ **Renewable resources** are those that essentially will never run out (are inexhaustible) – for example, wind, tides and direct solar energy.
- ◆ **Potentially renewable resources** can be renewed, but they will run out if we use them more quickly than they can be renewed. Examples include fresh water and air, fertile soil, and plant and animal biomass.

The biggest environmental concern is the depletion of non-renewable resources. Once they are used up, we will have to manage without them. Metal ores, especially those of iron, aluminium and copper, are becoming scarcer. The ores that still exist are often of low quality, making the process of extraction costly. Fossil fuels are another concern. New deposits of oil are being discovered but the speed at which we are using the oil we have is increasing. A time will come when all the oil, and eventually all the coal, will run out. Phosphate minerals, essential for the manufacture of fertilisers, are also becoming scarcer.

A number of these problems can be reduced by recycling some of the substances we use: recycling metals helps conserve metal ores and recycling plastics helps conserve the petroleum from which they are made. All recycling helps save energy, which comes mainly from fossil fuels.

Fossil fuels are a bigger problem. We will always need energy. A partial solution is to make more use of our renewable resources. Wind power, solar power and water power from rivers, tides and waves can all be used to generate electricity.

An increasing problem is the way in which our potentially renewable resources are being affected by overuse and pollution. The next three sections give more detail on these problems.

Questions

- 1.1 How does the Sun keep the water cycle working?
- 1.2 How does the Sun keep the carbon cycle working?
- 1.3 How could the Sun be the source of our energy in the future?
- 1.4 Write the word equations for:
 - a photosynthesis
 - b the complete combustion of carbon in air
 - c respiration.

S

1.2 The atmosphere

Uses of the gases of the air

Clean air has the following approximate composition: nitrogen 78%, oxygen 21%, argon 0.9% and 'other gases' (including carbon dioxide, water vapour, neon and other **noble gases**) 0.1% (Figure 1.6, overleaf).

Carbon dioxide is an important part of the air but makes up only about 0.04% of it. The carbon dioxide which is used by humans is not usually obtained from the air.

Nitrogen is used in the manufacture of ammonia and fertilisers in the **Haber process**. Liquid nitrogen is used in cryogenics (the storing of embryos and other types of living tissue at very low temperatures). Nitrogen is also sometimes used where an unreactive gas is needed to keep air away from certain products; for example, it is used to fill bags of crisps (chips) to ensure that the crisps do not get crushed or go rancid as a result of contact with oxygen in the air.

The biggest single use of oxygen is in the production of steel from cast iron. It is also used in oxyacetylene torches to produce the high-temperature flames needed to cut and weld metals. In hospitals, oxygen in cylinders is used to help the breathing of sick people.

Activity 1.1 Estimating the amount of oxygen in air

This is a demonstration of the reduction in volume when air is passed over heated copper.

A worksheet is included on the CD-ROM.



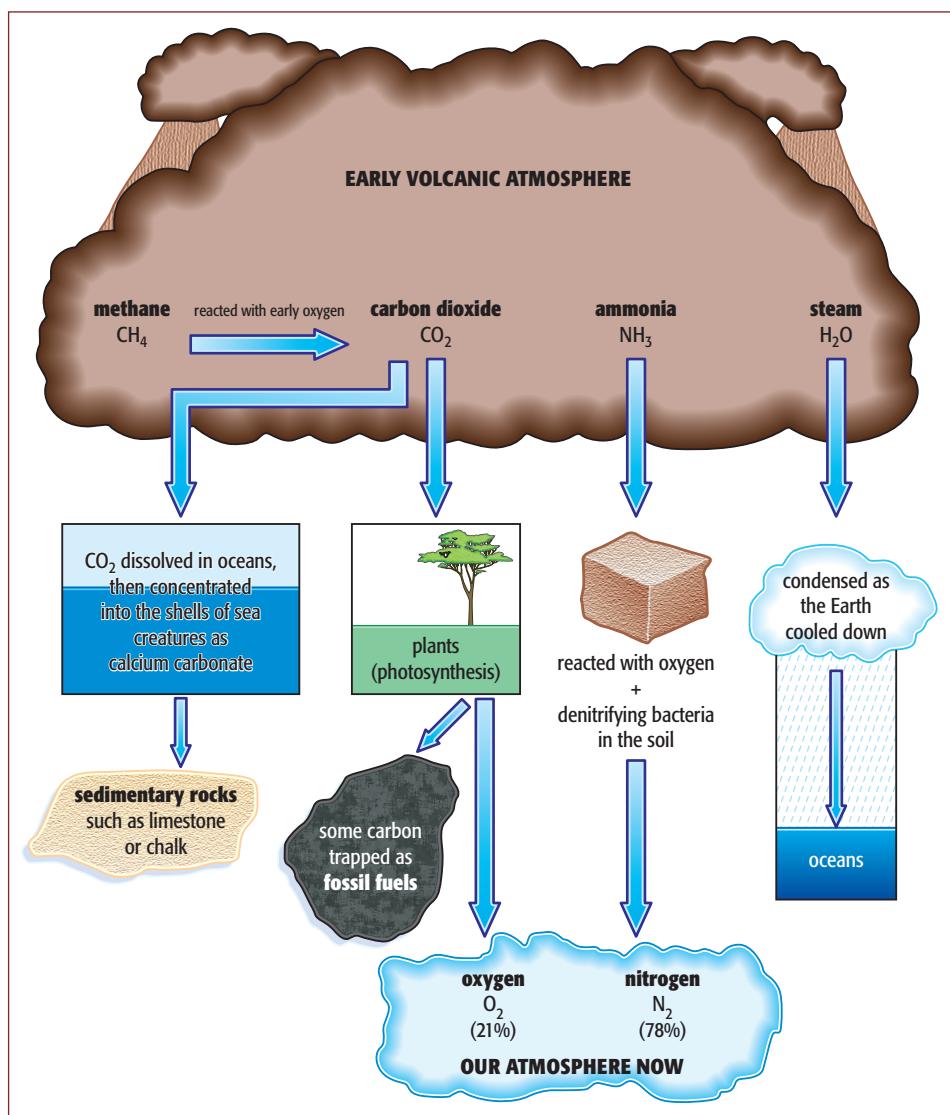


Figure 1.6 The development of the Earth's atmosphere.

Study tip

If you are asked for a use of oxygen, 'breathing' is not considered to be a correct answer because it is air rather than oxygen that we breathe. You need to give a use of pure oxygen.

because they glow with different colours when electricity flows through them.

Before any of the gases in the air can be used separately, they have to be separated from the air in the atmosphere. The method used is fractional distillation, which works because the gases have different boiling points (Table 1.1).

S

Study tip

Remember to be careful with temperatures below 0°C (with a negative sign). The boiling point of nitrogen (-196°C) is a lower temperature than -183°C (the boiling point of oxygen).

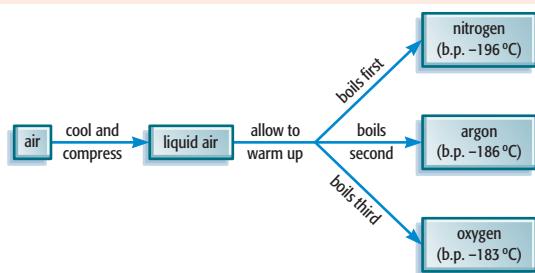
Argon and other noble gases are used in different types of lighting. Argon is used to 'fill' light bulbs to prevent the tungsten filament burning away (Figure 1.7). It does not react with tungsten even at very high temperatures. The other noble gases are used in advertising signs



Figure 1.7 Filament light bulbs contain argon which does not react with the hot tungsten filament.

The process of **fractional distillation** involves two stages.

- ◆ First the air must be cooled until it turns into a liquid (liquefies).
- ◆ Then the liquid air is allowed to warm up again. The various gases boil off one at a time at different temperatures.



Pollution of the air

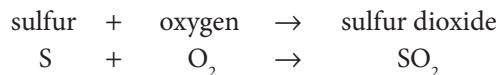
Many gases are accidentally or deliberately released into the air. Some are harmless but many create problems for the environment. The main source of ‘problem’ gases is the burning of fossil fuels.

Gas	Boiling point / °C	Proportion in mixture / %
carbon dioxide (sublimes)	-32	0.04
xenon	-108	— ^(a)
krypton	-153	— ^(a)
oxygen	-183	21
argon	-186	0.9
nitrogen	-196	78
neon	-246	— ^(a)
helium	-249	— ^(a)

^(a)All the other gases in the air make up 0.06% of the total.

Table 1.1 The boiling points of the gases in air.

Most countries produce electricity by burning coal or oil. Both these fuels are contaminated with sulfur, which produces sulfur dioxide when it burns:



Oxides of nitrogen (NO_x) (for example, nitrogen dioxide, NO_2) are also produced when air is heated in furnaces. These gases dissolve in rainwater to produce ‘acid rain’ (Figure 1.8, overleaf).

There are numerous effects of **acid rain**.

- ◆ Limestone buildings, statues, etc., are worn away.
- ◆ Lakes are acidified, and metal ions (for example, Al^{3+} ions) that are leached (washed) out of the soil damage the gills of fish, which may die.
- ◆ Nutrients are leached out of the soil and from leaves. Trees are deprived of these nutrients. Aluminium ions are freed from clays as aluminium sulfate, which damages tree roots. The tree is unable to draw up enough water through the damaged roots, and it dies.

The wind can carry acid rain clouds away from the industrialised areas, causing the **pollution** to fall on other countries.

One way to remedy the effects of acid rain is to add lime to lakes and the surrounding land to decrease the acidity. The best solution, however, is to prevent

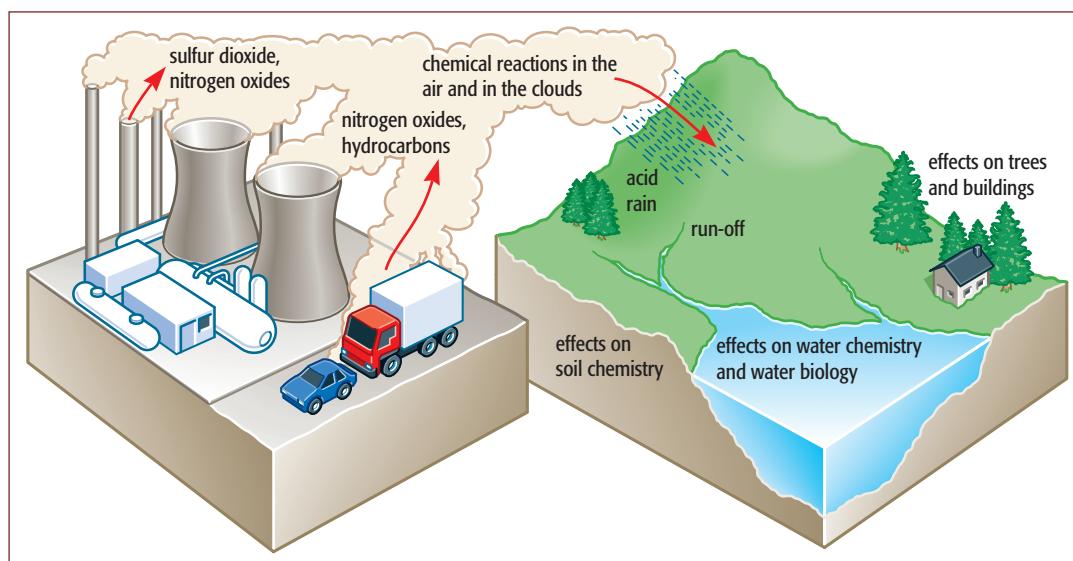
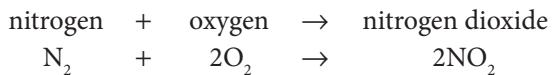


Figure 1.8 The formation of acid rain.

the acidic gases from being released in the first place. ‘Scrubbers’ are fitted to power station furnaces. In these devices, the acidic gases are passed through an alkaline substance such as lime. This removes the acids, making the escaping gases much less harmful. In many countries, though, acidic gases from power stations are still a serious problem.

Petrol (gasoline) and diesel for use in road transport have most of their sulfur removed when they are refined. Sulfur dioxide is not a serious problem with motor vehicles but the other contents of vehicle exhaust fumes (Figure 1.9) can cause problems. Nitrogen dioxide, for example, is still produced. The high temperature inside the engine’s cylinders causes the nitrogen and oxygen in the air to react together:



Because of the lack of oxygen in the enclosed space of an engine, the fuel does not usually burn completely and carbon monoxide (CO) is formed.



Figure 1.9 Fumes from a car exhaust.

Another pollution problem arising from motor vehicles is caused by tetraethyl lead in petrol (leaded petrol). Burning this type of petrol releases the toxic metal lead into the environment (Figure 1.9). The use of lead in petrol has decreased significantly over the last 20 years. In 2011, the United Nations announced the successful, worldwide, phasing out of leaded petrol for road vehicles. There are only a handful of countries where it is still available.

The dangers of these pollutants are as follows.

- ◆ **Nitrogen dioxide** causes acid rain and can combine with other gases in very hot weather to cause **photochemical smog**. This contains low-level ozone and is likely to cause breathing problems, especially in people with asthma.
- ◆ **Carbon monoxide** is a highly toxic gas. It combines with the haemoglobin in blood and stops it from carrying oxygen. Even very small amounts of carbon monoxide can cause dizziness and headaches. Larger quantities cause death.
- ◆ **Lead** is a neurotoxic metal and can cause learning difficulties in children, even in small quantities. The body cannot easily get rid of lead, so small amounts can build up to dangerous levels over time.

There are solutions to some of these problems.

Catalytic converters can be attached to the exhaust systems of cars (Figure 1.10). These convert carbon monoxide and nitrogen dioxide into carbon dioxide and nitrogen. Unfortunately, if there is lead in the petrol being used, the **catalyst** becomes poisoned and will no longer work. This means that in countries

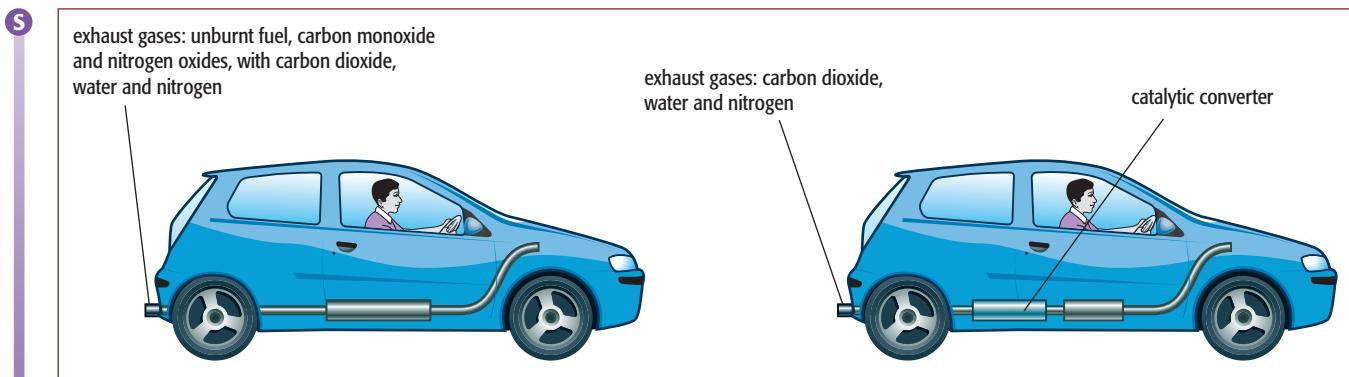


Figure 1.10 A catalytic converter changes harmful exhaust gases into safer gases.

where leaded petrol is still being used, catalytic converters cannot be used either.

Study tip

Try to keep these different atmospheric pollution problems clear and distinct in your mind rather than letting them merge together into one (confused?) problem. They each have distinct causes and clear consequences.

Figure 1.11 summarises the effects of the main pollutants of the air.

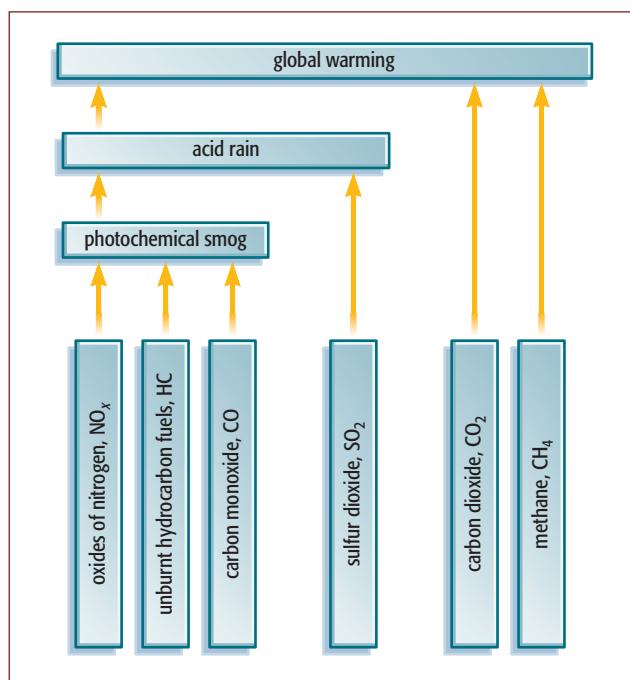


Figure 1.11 A summary of various atmospheric pollution problems caused by human activity.

Global warming and the 'greenhouse effect'

There are two gases in Figure 1.11, carbon dioxide and methane, which are not in the list of pollutants given so far. These, together with water vapour and oxides of nitrogen, are causing **global warming** due to the '**greenhouse effect**'. The Earth is warmed by the Sun but this heat would quickly escape if it were not for our atmosphere. It is always colder on a clear night because there are no clouds to keep the heat in. Some gases are better at keeping heat in than others; if there is too much of these gases in the atmosphere, the Earth gets warmer and this causes problems (Figure 1.12).

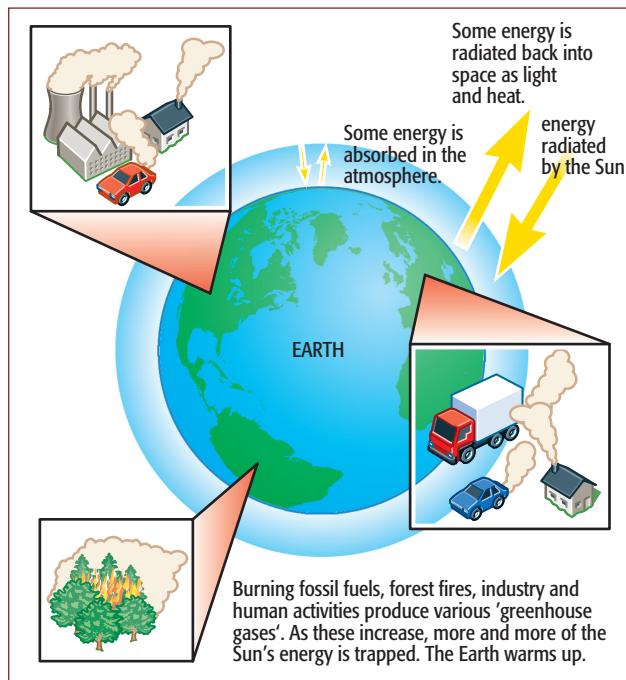


Figure 1.12 The greenhouse effect.

Some of the problems global warming will cause are listed below.

- ◆ Glaciers and polar ice will melt. This will cause a rise in sea level, and low-lying land will be flooded.
- ◆ The surface temperature of the Earth will increase. Deserts will spread and millions of people will have less water.
- ◆ Severe weather events will increase in frequency, and hurricanes and flooding will become more common.
- ◆ In some areas it may become easier to grow food crops but in others it will certainly become more difficult.

Carbon dioxide and methane are the two main problem gases; methane is around 20 times more effective at stopping heat escaping than carbon dioxide is.

Carbon dioxide enters the air through respiration and burning and it is removed by plants during photosynthesis. Burning more fuel and cutting down the forests increase the problem. Burning less fossil fuel and planting more trees would help to solve it.

Methane is produced by animals such as cows: it is a by-product of digesting their food. It emerges from both ends of the cow (but mostly from the mouth). Intriguingly, termites are also significant contributors to the methane in the atmosphere (Figure 1.13a). In addition, it is produced by the decay of food and other dead organic matter. It is produced in large quantities by rice paddy fields (Figure 1.13b) and landfill sites. Treating

organic waste so that the methane could be collected and burnt as fuel would help solve the problem.

The warming of the Arctic region in recent years has heightened our awareness of a further source of methane, known to scientists as ‘fire ice’ because it can ignite spontaneously. Melting of the Arctic ice and the consequent release of the large amount of the gas stored in the permafrost could have a huge economic and damaging environmental impact.

Questions

- 1.5 Which gases contribute most significantly to acid rain?
- 1.6 How do the gases responsible for acid rain get into the atmosphere?
- 1.7 What are the problems caused by acid rain?
- 1.8 What is photochemical smog and why is it a problem?
- 1.9 How does carbon monoxide stop the blood from carrying oxygen?
- 1.10 Why are light bulbs filled with argon?
- 1.11 How does methane get into the air?
- 1.12 What is the ‘greenhouse effect’?
- 1.13 What does a catalytic converter do to the exhaust gases from a car?
- 1.14 Why is it possible to separate the gases in the air by fractional distillation?



Figure 1.13 **a** A termite mound in Northern Territory, Australia – termites produce methane from digestion in their guts. **b** Terraced rice fields in Bali, Indonesia. Rice is the staple diet of about half the world’s population.

S

1.3 Seas and rivers

There is plenty of water on the Earth but most of it is in seas and oceans and the salts dissolved in it make it unsuitable for many uses. The amount of fresh water (less than 3% of the total) is still sufficient but it is not always in the places where it is needed. Figure 1.14 shows how the Earth's water is distributed.

You will see from the diagram that less than a teaspoon of water out of every 100 dm^3 is easily available for human use. This would be enough but it is not equally distributed around the world: rainforests can have more than 11 metres of rainfall in a year and desert areas less than one centimetre. There are places on Earth where it hasn't rained for more than ten years.

Water is essential to life but it can also carry disease. Polluted water kills many millions of people every year. It is important that the water we drink is treated to make it safe, and even more important that sewage (human and animal waste) is treated before being allowed back into rivers used for drinking water.

Water treatment

Water from rivers and lakes, and from underground, can contain dissolved salts, solid particles and bacteria.

The water purification process is designed to remove the last two of these. At its simplest, water treatment involves filtering the water to remove solid particles and adding chlorine to kill any bacteria that could cause disease.

Figure 1.15 (page 13) shows a modern water treatment process. The main difference from the simple treatment is in the use of ozone to remove pesticides and some other dissolved substances which can cause health problems. The water is still not totally pure as it contains some dissolved solids. Some of these, such as calcium salts, can aid health, whereas others, such as nitrate fertilisers, can be harmful.

In some parts of the world, seawater is made drinkable by **desalination** (taking the salt out). This can be done by distillation or by forcing the water through special membranes using high pressures (reverse osmosis). Desalination is particularly important in countries such as Saudi Arabia.

Our water supply is very important. Not only is it used in the home, as shown in Figure 1.16 (page 13), but it is also used in large quantities by industry. Most of the water used by industry is utilised as a solvent for other substances, to cool down reactions or to transfer heat from one part of a factory to another.

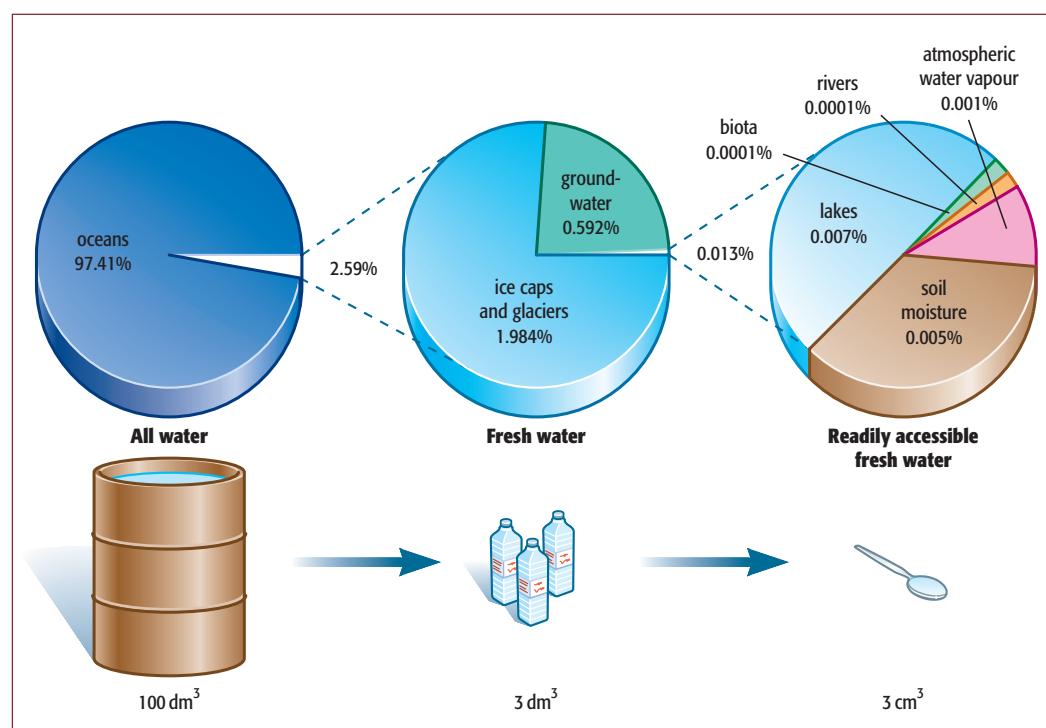


Figure 1.14 The availability of water on Earth.



Activity 1.2

Chemicals from seawater

Skills

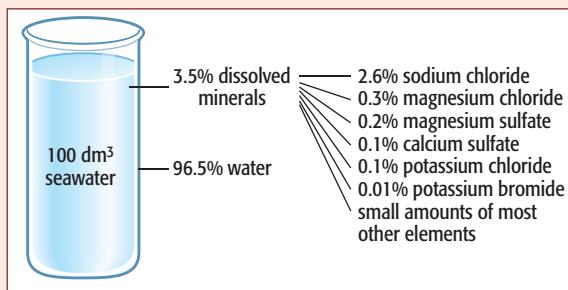
- A03.1 Using techniques, apparatus and materials
- A03.3 Observing, measuring and recording
- A03.4 Interpreting and evaluating observations and data



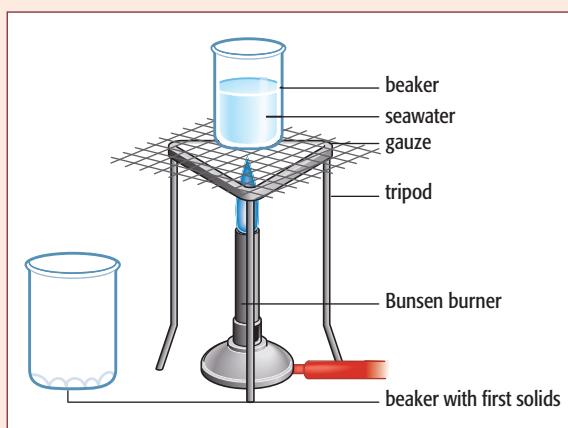
Wear eye protection throughout.

Take care with hot apparatus and solutions.

The sea is mainly water but there are lots of other things in it too. The most common substance in seawater is sodium chloride, or common salt. Other substances in it include calcium sulfate, magnesium sulfate and tiny amounts of metals such as copper and iron.



This experiment is designed to show that seawater contains a mixture of different salts.



- 1 Place 200 cm³ of seawater in a 250 cm³ beaker.
- 2 Heat and boil the seawater.
- 3 Stop heating when about 60–70 cm³ of liquid remains. Solid will be precipitated during this evaporation process.

- 4 Allow to cool and let any solids settle.
- 5 Pour the clear liquid into a 100 cm³ beaker, leaving the solids behind.
- 6 Add a few drops of dilute hydrochloric acid to the solids left behind and observe what happens.
- 7 Put the 100 cm³ beaker on the tripod and gauze and heat the liquid until another solid appears. This will occur when about 30–40 cm³ of liquid remains.
- 8 Carefully filter the liquid into a conical flask.
- 9 Wash out the 100 cm³ beaker and pour the filtrate into the beaker.
- 10 Boil the liquid again until there is almost none left.
- 11 Let it cool and note what you observe.

The role of the oceans in the carbon cycle: exchanging carbon dioxide between the atmosphere and ocean

- 1 Pour 100 cm³ of seawater into one beaker and 100 cm³ of fresh (tap) water into another.
- 2 Add several drops of Universal Indicator to each so that the colour is clearly visible.
- 3 Next, using a straw, blow gently and consistently into the water samples – first the seawater, then the fresh water. In each case, time how long it takes the indicator to become yellow. Record the results.



Questions

- A1 What evidence is there that seawater is a mixture of salts?
- A2 What gas is likely to have been given off when hydrochloric acid is added to the solids first collected?
- A3 What does this tell you about the identity of these solids?
- A4 Search the internet to try to find information about the solubilities of sodium chloride and calcium sulfate – two common compounds present in seawater. Use this information to predict the possible identity of the final solid left at the end of your experiment.

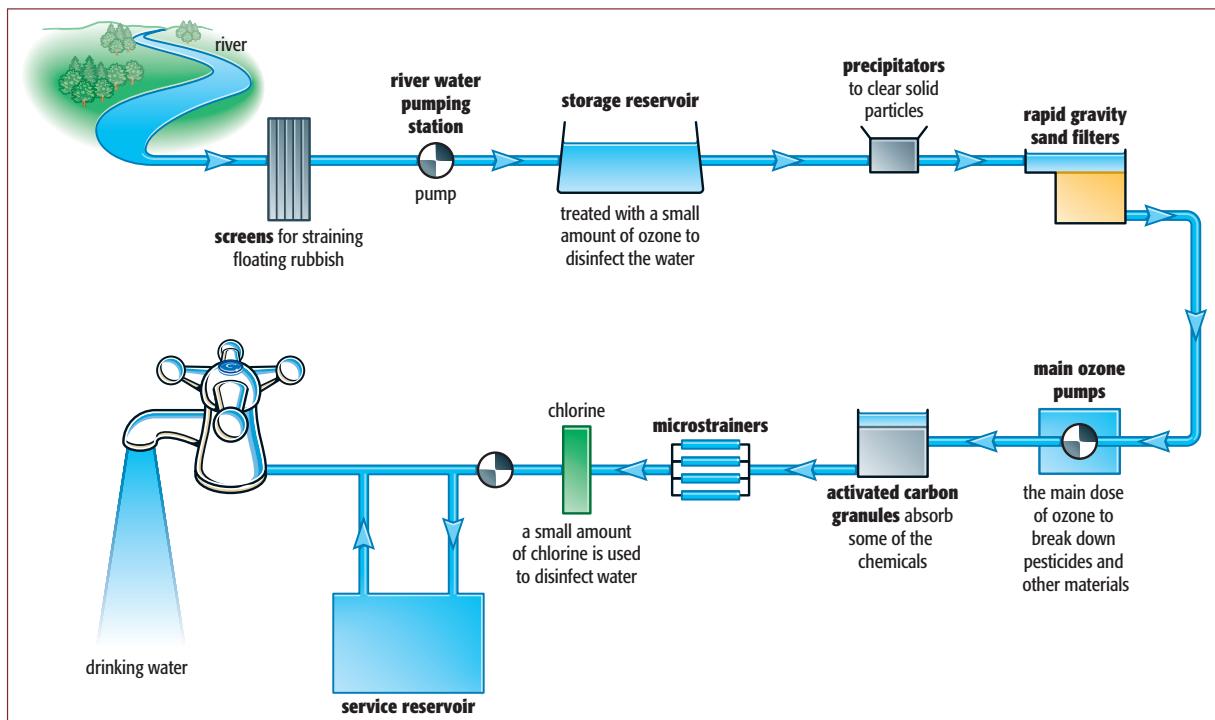


Figure 1.15 Purifying water for the domestic and industrial supply.

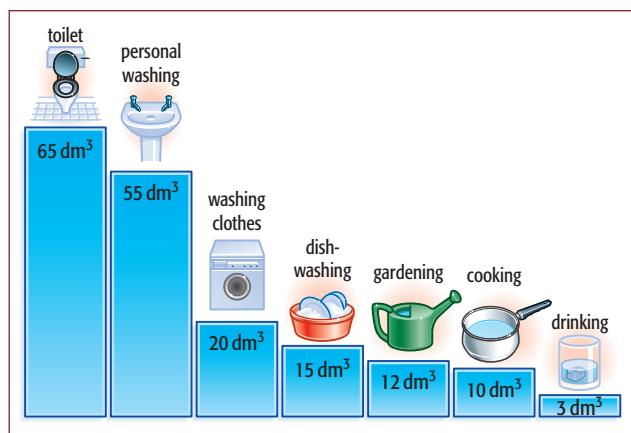


Figure 1.16 The main uses of water in a UK home. The numbers show how much water is used on average per person for each activity every day.

Pollution of the water supply

Issues concerning the pollution of water include the following:

- ◆ Nitrate fertilisers can be washed into streams and rivers from farmland. These nitrates are not removed by water treatment and can cause health problems for old people and for young children. They also cause waterways to become overloaded with nutrients, causing plant and algal growth.

- ◆ Industry sometimes discharges toxic and harmful substances into rivers.
- ◆ Untreated **sewage** and other animal waste can be released into rivers, especially in areas where there are no sewers. Release of untreated sewage into rivers can lower oxygen levels by chemical reaction and cause the spread of harmful bacteria, increasing the risk of disease.
- ◆ The use of water for cooling by industry can result in warm water being discharged into rivers. Warm water can dissolve less oxygen than cold water, so animals living in the water may be left with insufficient oxygen. All of these, and others, lead to problems in rivers and lakes.

Questions

- 1.15 Why is water filtered before other treatments?
- 1.16 Why is chlorine added to water?
- 1.17 Why is distillation of seawater an expensive way of making drinking water?
- 1.18 What is the main danger of letting untreated sewage into rivers?
- 1.19 Why are nitrates from fertilisers dangerous in the water supply?

1.4 The Earth's crust

The Earth's **crust** is the top layer of solid rock of the planet. As the Earth's crust moves, rock is constantly being taken down into the molten rock beneath the surface. This rock is changed and sometimes decomposed before it rises back to the surface and cools. These processes give rise to different types of rock and we extract some of those near the surface by mining and quarrying for human use. The decomposition also produces gases, mainly carbon dioxide and water vapour, which, together with molten rock, still escape from the Earth's **crust** through volcanoes.

This **rock cycle** is powered by energy produced by radioactive decay and heat from the Earth's **core**. It is a very slow process – the plates of the Earth's crust are moving only a few centimetres each year.

The crust varies in thickness from 5 km below some parts of the ocean to around 50 km in some parts of the land mass. Since the distance from the Earth's surface to its **core** is over 6000 km, the crust is a very thin surface layer. The crust is where the majority of the chemicals that we use come from.

Metal **ores** are rocks that have a relatively high concentration of a mineral containing a certain metal. For more details of ores and methods of obtaining metals from them, see Chapters 8 and 9.

Rocks can be used for building and for the extraction of useful chemicals other than metals. The most useful of these is **limestone**.

The limestone cycle

Limestone is an important resource from which a useful range of compounds can be made. Figure 1.17 shows some of the important uses of limestone and the related compounds quicklime and slaked lime. The reactions involved in producing these compounds can be imitated in the laboratory (Figure 1.18).

A piece of calcium carbonate can be heated strongly for some time to produce **lime** (quicklime, calcium oxide). The piece of lime is allowed to cool and then a few drops of water are added. The solid flakes and expands, crumbling into 'slaked lime'. This reaction is strongly **exothermic**. If more water is added, an alkaline solution (limewater) is obtained. The cycle can be completed by bubbling carbon dioxide into the solution. A white precipitate of calcium carbonate is formed. We can complete what is sometimes referred to as the **limestone cycle** (Figure 1.19).

Activity 1.3 Thermal decomposition of calcium carbonate

This activity illustrates some of the chemistry of limestone (calcium carbonate) and other materials made from it. The experiment demonstrates the 'limestone cycle'.

A worksheet is included on the CD-ROM.

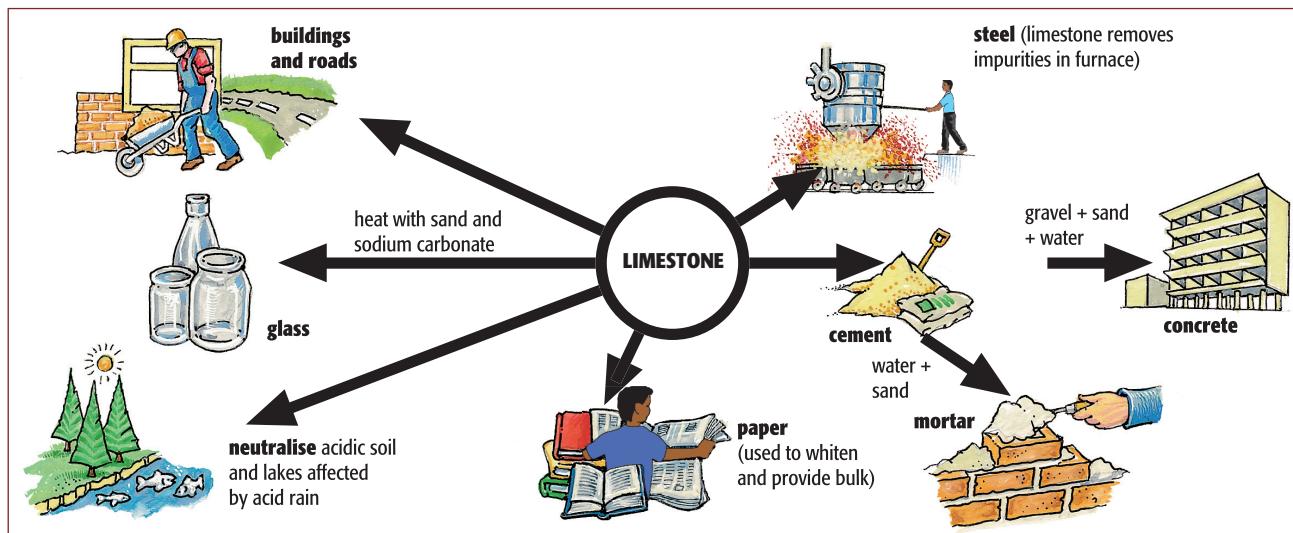


Figure 1.17 Some of the uses of limestone (calcium carbonate).

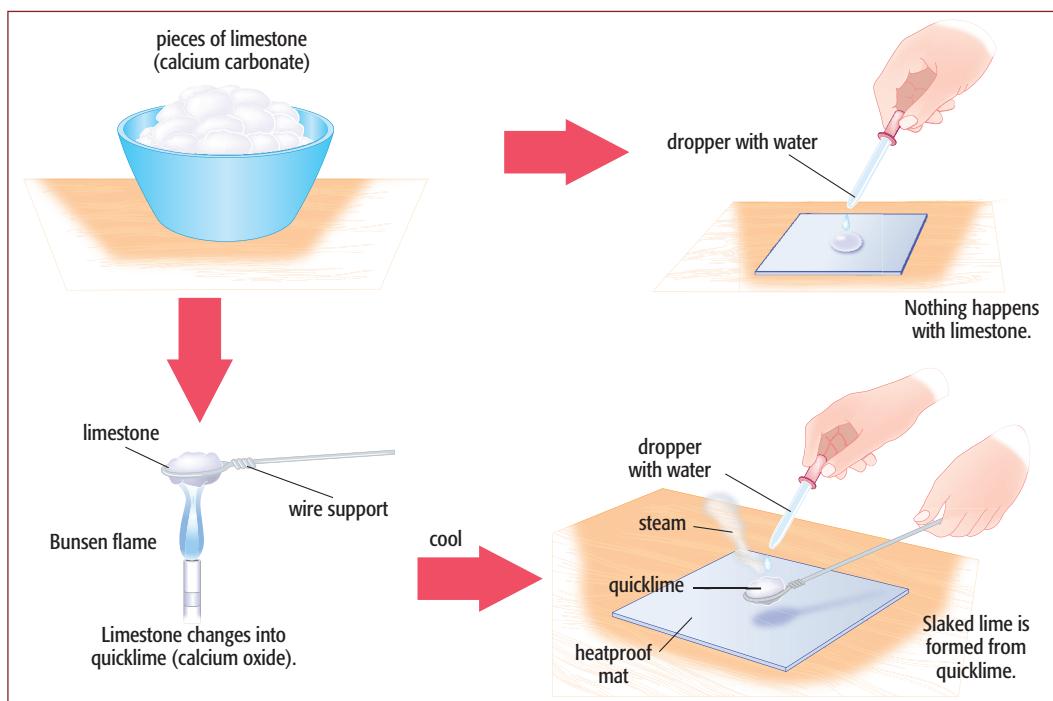


Figure 1.18 The formation of quicklime and slaked lime in the laboratory.

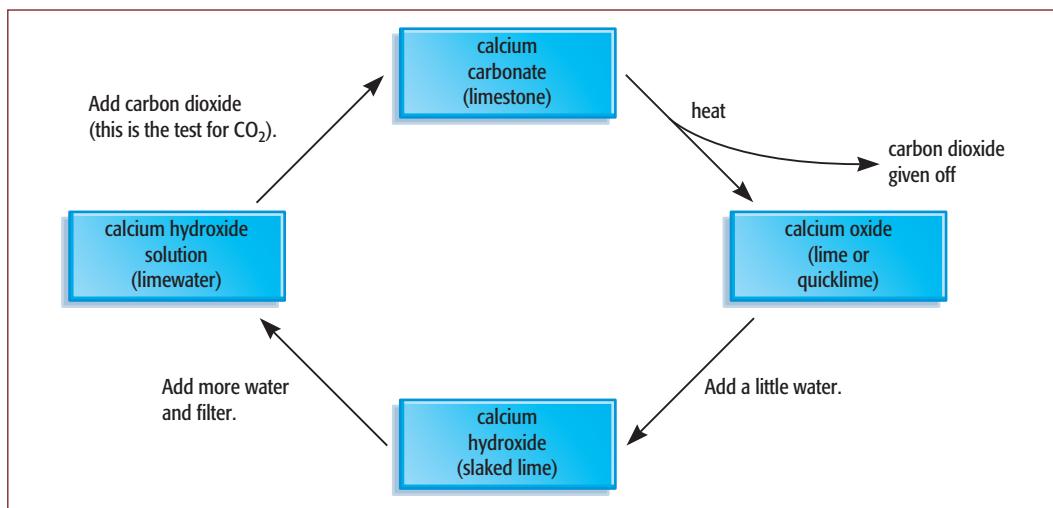


Figure 1.19 The limestone cycle.

More detail on the importance of limestone and the chemicals derived from it can be found in Section 9.6. This includes the method of making lime industrially.

The problem of fossil fuels

The major **fossil fuels** are **coal**, **petroleum** (or **crude oil**) and **natural gas**. These are important sources of energy but are also very important as sources of raw materials for making plastics, drugs, detergents and

many other useful substances. For more detail of fossil fuels and their uses see Chapter 11.

If these resources were used not as fuels but only as a source of the chemicals we need, there wouldn't be a problem. They are, however, mostly used as fuels and they are a limited, non-renewable resource.

In 2007, it was estimated that the fossil fuel supplies we currently know about would last a further 43 years

in the case of petroleum, 167 years for natural gas and 417 years for coal.

As more countries become industrialised, energy use in the world is increasing at an even faster rate than the population. Table 1.2 shows the sources of energy drawn on in the years 2006 and 2011.

The figures show how relatively little energy is obtained from sources other than fossil fuels, with the figure for coal showing a notable increase between 2006 and 2011. Hydroelectric power (6.4%) and nuclear energy (4.9%) do make a contribution, although the figure for nuclear energy has fallen in this time. In 2011, only 1.6% of energy came from solar power, wind power and wave power combined.

Clearly change is necessary (Figure 1.20). The Sun is the greatest provider of energy to the Earth. The amount of solar energy falling on the Earth's surface is immensely vast. In one year it is about twice as much as will ever be obtained from all of the the Earth's non-renewable resources combined.

Energy source	Percentage of energy from this source / %	
	for 2006	for 2011
petroleum	36.8	33.1
coal	26.6	30.3
natural gas	22.9	23.7
hydroelectric power	6.3	6.4
nuclear energy	6.0	4.9
renewables (solar, wind, etc.)	1.4	1.6

Table 1.2 The sources of energy used worldwide in 2006 and 2011.

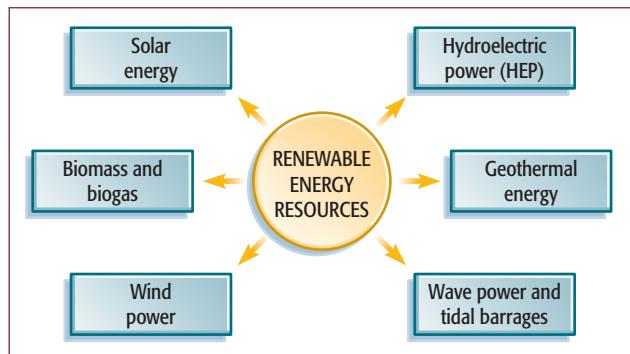


Figure 1.20 Renewable energy resources suitable for development to reduce our dependence on fossil fuels.

Hydrogen is one possible fuel for the future, either as a substance to burn or for use in fuel cells.

Hydrogen as a fuel

Hydrogen gas has attractions as a **fuel**. All it produces on burning is water. When hydrogen burns, it produces more energy per gram than any other fuel (Figure 1.21).

A future 'hydrogen economy' has been talked about, but there are problems of storage and transport. The gas itself is difficult to store and transport because of its low density. The first vehicles to run on hydrogen were the rockets of the US space programme. Hydrogen is not cheap. The main method of obtaining it on a large scale is by the electrolysis of water. However, this is not very economical. It is possible that cheap surplus electricity from nuclear power may make electrolysis more economical. Others have suggested the use of electricity from solar power.

Despite these difficulties, prototype hydrogen-powered cars have been tried. Nissan and Mazda in Japan, and BMW and Daimler-Benz in Germany, are among those who have built and tested cars. The Japanese prototype burns the hydrogen in the engine, while the German–Swiss–British venture uses the hydrogen in a **fuel cell**. Electricity from this cell then powers an electric motor (Figure 1.22). Using a fuel cell operating an electric motor, hydrogen has an

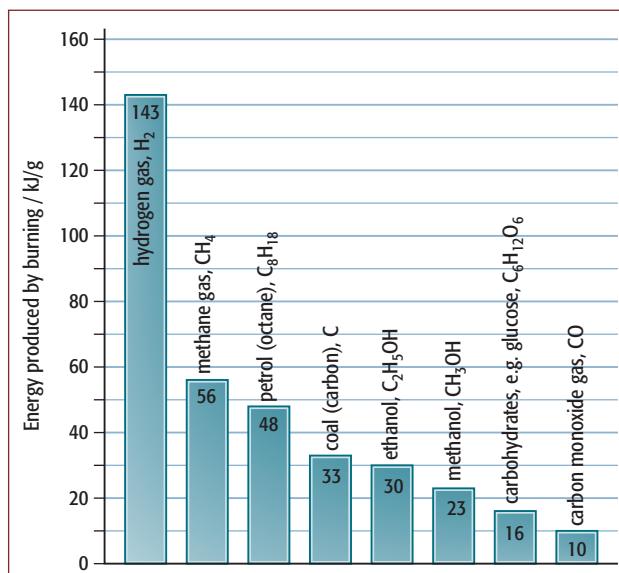


Figure 1.21 The energy produced on burning one gram of various fuels, to produce water and carbon dioxide. Hydrogen produces more energy per gram than any other fuel.

efficiency of 60% compared with 35% for a petrol engine. The ‘hydrogen economy’ may have life in it yet! The advantages and disadvantages are summarised in Table 1.3.

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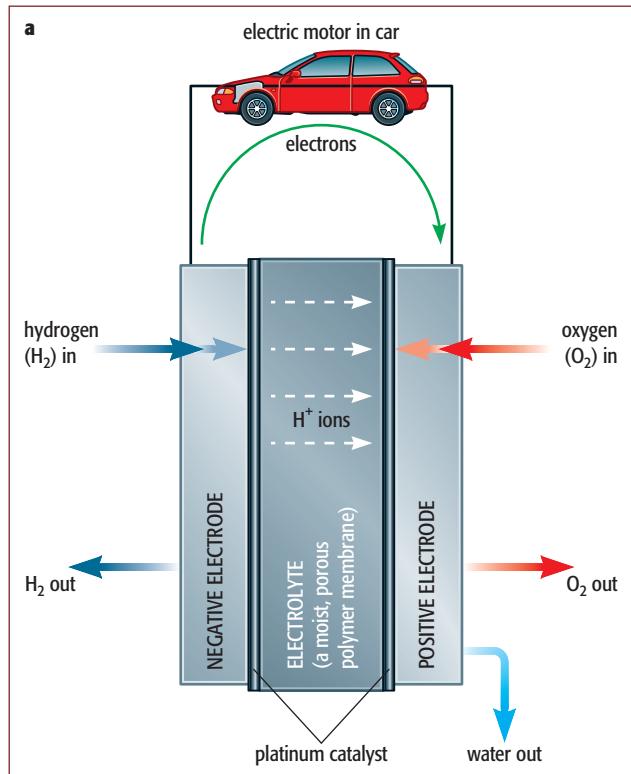
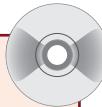


Figure 1.22 **a** How a car runs on a hydrogen fuel cell. The car is powered by electrons released at the negative electrode. Inside the fuel cell, hydrogen ions move to the positive electrode, where they react with oxygen to form water. **b** These hydrogen-powered minicabs are part of a development project at the University of Birmingham in the UK.

Advantages	Disadvantages
<ul style="list-style-type: none"> ◆ renewable if produced using solar energy ◆ lower flammability than gasoline (petrol) ◆ virtually emission-free ◆ zero emissions of CO₂ ◆ non-toxic 	<ul style="list-style-type: none"> ◆ non-renewable if generated using nuclear energy or energy from fossil fuels ◆ large fuel tank required ◆ as yet there are very few ‘filling stations’, where a car could be topped up with hydrogen ◆ engine redesign needed or a fuel cell system ◆ currently expensive

Table 1.3 The advantages and disadvantages of hydrogen as a fuel for motor vehicles.



Activity 1.4 Hydrogen power – communicating the benefits

Groups of students are to design a flyer, leaflet or web page to be used as publicity by a car dealership outlining the issues and benefits of hydrogen-powered cars. Use the internet to find information.

A worksheet is included on the CD-ROM.

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Activity 1.5 Using solar and hydrogen power

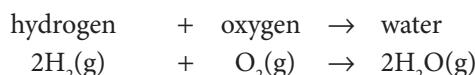
This demonstration connects a small solar panel to a proton exchange membrane (PEM) fuel cell. In this mode, the cell acts as an electrolyser and can decompose distilled water into its elements. The experiment can then be reversed and the gases collected can be used to power the fuel cell to drive a small electric fan.

A worksheet is included on the CD-ROM.

Hydrogen fuel cells

Research has found a much more efficient way of changing chemical energy into electrical energy by

S using a fuel cell. A hydrogen fuel cell can be used to power a car. Such a cell operates continuously, with no need for recharging. The cell supplies energy as long as the reactants are fed in to the electrodes. The overall reaction of the hydrogen–oxygen fuel cell is:



Land pollution

Cities throughout the world are covered in litter. Some make an effort to control it but it is always there. Most of our waste material is buried and this can lead to problems. Toxic and radioactive waste can make the land unusable and many countries strictly control what can be buried and where. Companies are required, by law, to treat their waste products to make them as harmless as possible.

Domestic waste should be recycled whenever possible. Waste that cannot be treated in this way should be burned to create energy. If it is left in landfill sites, it

decays, producing methane gas (by a process known as **anaerobic decay**). Methane is much more harmful to the environment than the carbon dioxide produced by burning it, as methane is a more powerful greenhouse gas.

Questions

- 1.20** What makes an ore different from any other type of rock?
- 1.21** What is the difference between lime and slaked lime?
- 1.22** What useful chemicals can be made from petroleum?
- 1.23** Name **four** types of ‘alternative energy’.
- 1.24** What makes a gas a ‘greenhouse gas’?
- 1.25** Give an advantage of using hydrogen as a fuel.
- 1.26** What is the essential reaction taking place in a hydrogen fuel cell? Give the word and balanced chemical equations for the reaction.

Summary

You should know:

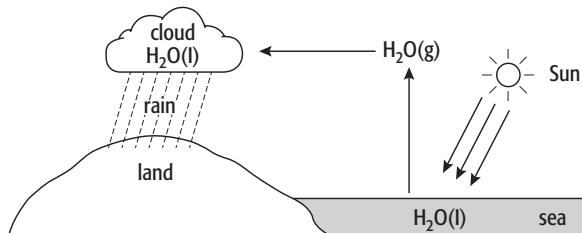
- ◆ that there are important natural cycles that involve the movement of resources within the Earth’s ecology, two of these being the water cycle and the nitrogen cycle
- S** ◆ that there is a third important natural cycle, the carbon cycle, which involves the key processes of photosynthesis and respiration
- ◆ that the air is composed predominantly of nitrogen and oxygen, but that other gases have major roles to play too
- ◆ about the major atmospheric pollution problems that are changing the nature of our world, including global climate change and acid rain
- ◆ that global warming is caused by an increase in the atmosphere of certain ‘greenhouse gases’ such as carbon dioxide and methane
- ◆ how the gases of the air can be separated by the fractional distillation of liquid air, and that the separated gases have their own uses
- ◆ how the availability of clean fresh water is one of the major problems in the world
- ◆ how mineral ores and crude oil provide sources of metals and chemicals for industrial use
- ◆ that limestone, one of these mineral resources, has a range of uses, from the making of cement and concrete to the extraction of iron in the blast furnace
- ◆ how we are currently highly dependent on fossil fuels as the source of our energy supply
- ◆ that hydrogen is one possible new energy source that is currently under development – it is seen as an environmentally clean fuel because the only product of its combustion is water
- S** ◆ how the hydrogen fuel cell is based on the production of electrical power using the combustion reaction for hydrogen and can be used to power cars.

End-of-chapter questions

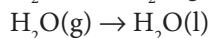
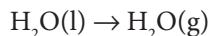
- 1 The water cycle and the carbon cycle are both of vital importance to life. Explain how they both depend on energy from the Sun.
- 2 Water is present in the atmosphere, in the seas and in ice and snow.
 - a Describe a chemical test for water. Give the test and the positive result. [2]
 - b State **one** use of water in industry. [1]
 - c Water is a good solvent. What do you understand by the term **solvent**? [1]
 - d Water vapour in the atmosphere reacts with sulfur dioxide, SO_2 , to produce acid rain.
 - i State **one** source of sulfur dioxide. [1]
 - ii State **two** adverse effects of acid rain. [2]
 - iii Calculate the relative molecular mass of sulfur dioxide. [1]
 - e Water from lakes and rivers can be treated to make the water safer to drink. Describe **two** of the steps in water purification. For each of these steps, give an explanation of its purpose. [4]
 - f Water is formed when hydrogen burns in air. State the percentage of oxygen present in the air. [1]

[Cambridge IGCSE® Chemistry 0620/21, Question 3(a-e & f(i)), June 2011]

- S** 3 The diagram below shows part of the water cycle.



- a i State the name of each of the following changes of state.



[2]

- ii Which **one** of the above changes of state is exothermic? Explain your choice. [1]
- b The rain drains into rivers and then into reservoirs. Describe how water is treated before it enters the water supply. [2]
- c Explain how acid rain is formed. [4]

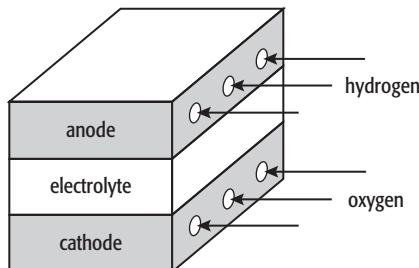
[Cambridge IGCSE® Chemistry 0620/31, Question 1(a, b & c(i)), June 2012]

- 4 Two important greenhouse gases are methane and carbon dioxide.

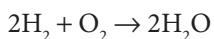
- a Methane is 20 times more effective as a greenhouse gas than carbon dioxide. The methane in the atmosphere comes from both natural and industrial sources.
 - i Describe **two** natural sources of methane. [2]
 - ii Although methane can persist in the atmosphere for up to 15 years, it is eventually removed by oxidation. What are the products of this oxidation? [2]
- b How do the processes of respiration, combustion and photosynthesis determine the percentage of carbon dioxide in the atmosphere? [4]

[Cambridge IGCSE® Chemistry 0620/31, Question 2, November 2011]

- S** 5 Fuel cells are used in spacecraft to produce electrical energy.



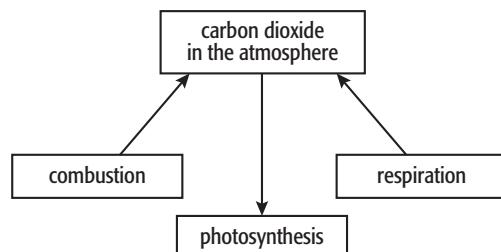
- a How is oxygen obtained from liquid air? [2]
 b Hydrogen and oxygen react to form water.



- i Give an example of bond breaking in the above reaction. [1]
 ii Give an example of bond forming in the above reaction. [1]
 iii Is the change given in part i exothermic or endothermic? [1]
 c i Give two reasons why hydrogen may be considered to be the ideal fuel for the future. [2]
 ii Suggest a reason why hydrogen is not widely used at the moment. [1]

[Cambridge IGCSE® Chemistry 0620/32, Question 5, June 2010]

- 6 The diagram shows part of the carbon cycle. This includes some of the processes that determine the percentage of carbon dioxide in the atmosphere.



- a Carbon dioxide is one greenhouse gas. Name another one. [1]
 b Explain the term **respiration** and how this process increases the percentage of carbon dioxide in the atmosphere. [3]
 c Explain why the combustion of waste crop material should not alter the percentage of carbon dioxide in the atmosphere. [2]
 d In 1960, the percentage of carbon dioxide in the atmosphere was 0.032% and in 2008 it was 0.038%. Suggest an explanation for this increase. [2]

[Cambridge IGCSE® Chemistry 0620/31, Question 7, November 2010]

2

The nature of matter

In this chapter, you will find out about:

- ◆ the three states of matter, and changes of state
- ◆ separating and purifying substances
- ◆ filtration
- ◆ use of a separating funnel
- ◆ evaporation and crystallisation
- ◆ distillation
- ◆ paper chromatography
- ◆ criteria of purity
- ◆ elements and compounds
- ◆ atomic theory
- ◆ the kinetic theory and changes of state
- ◆ diffusion
- ◆ Brownian motion
- ◆ atomic structure and subatomic particles
- ◆ proton (atomic) number and nucleon (mass) number
- ◆ isotopes
- ◆ relative atomic mass
- ◆ uses of radioactivity
- ◆ the arrangement of electrons in atoms.

Lord of the rings

Saturn is perhaps the most beautiful of the planets of the Solar System. It has fascinated astronomers because of its mysterious rings (Figure 2.1). The Pioneer, Voyager and Cassini–Huygens space-probes sent back a great deal of information on the nature of the rings and the mass of Saturn itself.

Each ring is made up of a stream of icy particles, following each other nose-to-tail around



Figure 2.1 Saturn and its rings silhouetted against the Sun: a photograph taken by the Cassini probe. The rings are made of ice and dust.

the planet. The particles can be of widely varying sizes. The rings resemble a snowstorm, in which tiny snowflakes are mixed with snowballs up to the size of a house. The ice that surrounds one of the most spectacular planets of our solar system is made of water – the same substance (with the same formula) that covers so much of the Earth's surface.

The planet of Saturn is made of gases, mainly hydrogen and helium. Deep in the centre of these lightweight gases is a small rocky core, surrounded by a liquid layer of the gases. The hydrogen is liquid because of the high pressure in the inner regions of the planet nearest the core. The liquid hydrogen behaves with metallic properties. Study of Saturn's physical structure emphasises how substances that we know on Earth can exist in unusual physical states in different environments.

How do changing conditions affect the appearance, properties and behaviour of different substances?

2.1 The states of matter

There are many different kinds of **matter**. The word is used to cover all the substances and materials of which the Universe is composed. Samples of all of these materials have two properties in common: they each occupy space (they have volume) and they have mass.

Chemistry is the study of how matter behaves, and of how one kind of substance can be changed into another. Whichever chemical substance we study, we find it can exist in three different forms (or physical states) depending on the conditions. These three different **states of matter** are known as **solid**, **liquid** and **gas**. Changing temperature and/or pressure can change the state in which a substance exists.

The different physical states have certain general characteristics that are true whatever chemical substance is being considered. These are summarised in Table 2.1.

Key definition

matter – anything that has mass and takes up space. There are three physical states: solid, liquid and gas.

The three physical states show differences in the way they respond to changes in temperature and pressure. All three show an increase in volume (an expansion) when the temperature is increased, and a decrease in volume (a contraction) when the temperature is lowered. The effect is much bigger for a gas than for either a solid or a liquid.

The volume of a gas at a fixed temperature can easily be reduced by increasing the pressure on the gas.

Gases are easy to ‘squash’ – they are easily compressed. Liquids, on the other hand, are only slightly compressible, and the volume of a solid is unaffected by changing the pressure.

Changes in physical state

Large changes in temperature and pressure can cause changes that are more dramatic than expansion or contraction. They can cause a substance to change its physical state. The changes between the three states of matter are shown in Figure 2.2. At atmospheric pressure, these changes can be brought about by raising or lowering the temperature of the substance.

Melting and freezing

The temperature at which a pure substance turns to a liquid is called the **melting point** (m.p.). This always happens at one particular temperature for each pure substance (Figure 2.3). The process is reversed at precisely the same temperature if a liquid is cooled down. It is then called the **freezing point** (f.p.). The melting point and freezing point of any given substance are both the same temperature. For example, the melting and freezing of pure water take place at 0 °C.

Gallium is a metal that has a melting point just above room temperature. Because of this it will melt in a person’s hand (Figure 2.4).

Sublimation

A few solids, such as carbon dioxide ('dry ice'), do not melt when they are heated at normal pressures. Instead, they turn directly into gas. This change of state is called **sublimation**: the solid **sublimes**. Like melting, this also happens at one particular temperature for each pure solid. Iodine is another solid that sublimes. It produces

Physical state	Volume	Density	Shape	Fluidity
solid	has a fixed volume	high	has a definite shape	does not flow
liquid	has a fixed volume	moderate to high	no definite shape – takes the shape of the container	generally flows easily ^(a)
gas	no fixed volume – expands to fill the container	low	no definite shape – takes the shape of the container	flows easily

^(a)Liquids and gases are called **fluids**.

Table 2.1 Differences in the properties of the three states of matter.

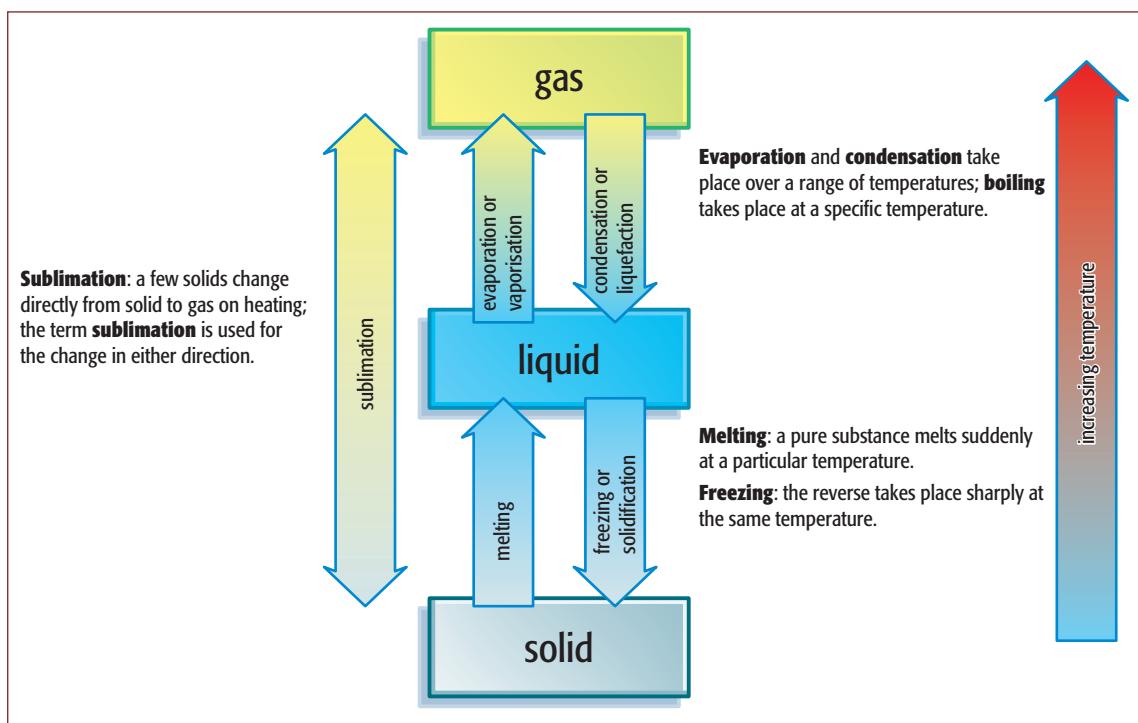


Figure 2.2 Changes of physical state and the effect of increasing temperature at atmospheric pressure.

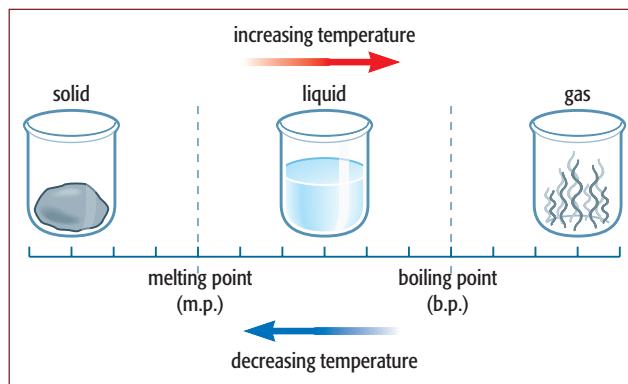


Figure 2.3 The relationship between the melting point and boiling point of a substance.

a purple vapour, but then condenses again on a cold surface (Figure 2.5, overleaf).

Evaporation, boiling and condensation

If a liquid is left with its surface exposed to the air, it evaporates. Splashes of water evaporate at room temperature. After rain, puddles dry up! When liquids change into gases in this way, the process is called **evaporation**. Evaporation takes place from the surface of the liquid. The larger the surface area, the faster the liquid evaporates.



Figure 2.4 The metal gallium has a melting point just above room temperature. It will literally melt in the hand.

The warmer the liquid is, the faster it evaporates. Eventually, at a certain temperature, it becomes hot enough for gas to form within the liquid and not just at the surface. Bubbles of gas appear inside the liquid. This process is



Figure 2.5 Iodine sublimes. On warming, it produces a purple vapour which then condenses again on the cool part of the tube.

known as **boiling**. It takes place at a specific temperature, known as the **boiling point** (b.p.) for each pure liquid (Figure 2.3). Water evaporates fairly easily and has a relatively low boiling point – it is quite a **volatile** liquid. Ethanol, with a boiling point of 78 °C, is more volatile than water. It has a higher **volatility** than water.

- ◆ A **volatile** liquid is one which evaporates easily and has a relatively low boiling point.
- ◆ Ethanol (b.p. 78 °C) is a more volatile liquid than water (b.p. 100 °C).

The reverse of evaporation is **condensation**. This is usually brought about by cooling. However, we saw earlier that the gas state is the one most affected by changes in pressure. It is possible, at normal temperatures, to condense a gas into a liquid by increasing the pressure, without cooling.

The boiling point of a liquid can change if the surrounding pressure changes. The value given for the boiling point is usually stated at the pressure of the atmosphere at sea level (**atmospheric pressure** or **standard pressure**). If the surrounding pressure falls,

the boiling point falls. The boiling point of water at standard pressure is 100 °C. On a high mountain it is lower than 100 °C. If the surrounding pressure is increased, the boiling point rises. In a pressure cooker, the boiling point of water is raised to around 120 °C and food cooks more quickly at this higher temperature.

Pure substances

A **pure substance** consists of only one substance. There is nothing else in it: it has no contaminating impurities. A pure substance melts and boils at definite temperatures. Table 2.2 shows the melting points and boiling points of some common substances at atmospheric pressure.

The values for the melting point and boiling point of a pure substance are precise and predictable. This means that we can use them to test the purity of a sample. They can also be used to check the identity of an unknown substance. The melting point can be measured using an electrically heated melting-point apparatus (Figure 2.6). A capillary tube is filled with a small amount of the solid and is placed in the heating block. The melting is viewed through a magnifying lens.

A substance's melting and boiling points in relation to room temperature (taken as 20 °C) determine whether it

Substance	Physical state at room temperature (20 °C)	Melting point/°C	Boiling point/°C
oxygen	gas	-219	-183
nitrogen	gas	-210	-196
ethanol (alcohol)	liquid	-117	78
water	liquid	0	100
sulfur	solid	115	444
common salt (sodium chloride)	solid	801	1465
copper	solid	1083	2600
carbon dioxide	gas	-78 ^(a)	

^(a)Sublimes.

Table 2.2 The melting and boiling points of some common chemical substances.



Figure 2.6 An electrical melting-point apparatus.

is usually seen as a solid, a liquid or a gas. For example, if the m.p. is below 20 °C and the b.p. is above 20 °C, the substance will be a liquid at room temperature.

Study tip

Remember to practise using melting and boiling point data to decide whether a particular substance is a solid, a liquid or a gas at room temperature. These are quite common questions.

Be careful with temperatures below 0 °C; –100 °C is a higher temperature than –150 °C.

The effect of impurities

Seawater is impure water. You can show this if you put some seawater in an evaporating dish and boil away the water, because a solid residue of salt is left behind in the dish. Seawater freezes at a temperature well below the freezing point of pure water (0 °C) and boils at a temperature above the boiling point of pure water (100 °C). Other impure substances show similar differences.

In addition, the impurity can also reduce the ‘sharpness’ of the melting or boiling point. An impure substance sometimes melts or boils over a range of temperatures, not at a particular point.

The presence of an impurity in a substance:

- ◆ lowers the melting point, and
- ◆ raises the boiling point of the substance.

Study tip

Remember that pure substances have definite, sharp melting and boiling points. The presence of an impurity means that these changes will be spread over a range of temperatures in each case.

Heating and cooling curves

The melting point of a solid can also be measured using the apparatus shown in Figure 2.7. A powdered solid is put in a narrow melting-point tube so that it can be heated easily. An oil bath is used so that melting points above 100 °C can be measured. We can follow the temperature of the sample before and after melting. These results can then be used to produce a heating curve (Figure 2.8, overleaf). Similar apparatus can be used to produce a cooling curve but the thermometer must be placed in a test tube containing the solid being studied.

Figure 2.8 shows how the temperature changes when a sample of solid naphthalene (a single pure substance)

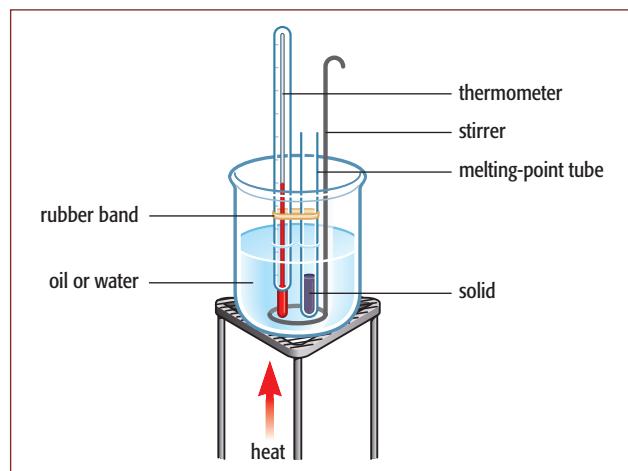


Figure 2.7 Apparatus for measuring the melting point of a solid. A water bath can be used for melting points below 100 °C and an oil bath for those above 100 °C.

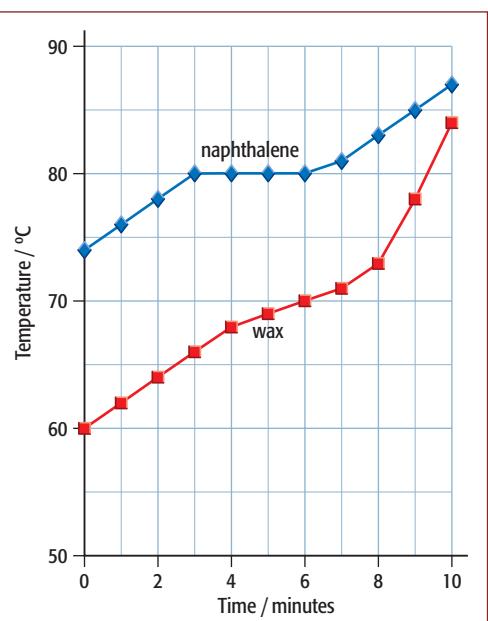


Figure 2.8 The heating curves for naphthalene (a pure substance) and wax (a mixture of substances).

is heated steadily. The solid melts at precisely 80 °C. Notice that, while the solid is melting, the temperature stops rising. It will only begin to rise again when all the naphthalene has melted. Generally, the heating curve for a pure solid stops rising at its melting point. The heating curve for wax, which is a mixture of substances, shows the solid wax melting over a range of temperatures.

It is possible to heat a liquid in the same apparatus until its boiling point is reached. Again, the temperature stays the same until all the liquid has boiled. The reverse processes can be shown if a sample of gas is allowed to cool. This produces a cooling curve (Figure 2.9). The level portions of the curve occur

where the gas condenses to a liquid, and when the liquid freezes.

These experiments show that heat energy is needed to change a solid into a liquid, or a liquid into a gas. During the reverse processes, heat energy is given out.

When a solid is melted, or a liquid is boiled, the temperature stays constant until the process is complete. The same is true in reverse when a gas condenses or a liquid freezes.



Activity 2.1 Plotting a cooling curve

In this experiment, you will plot cooling curves for two different substances.

A worksheet, with a self-assessment checklist, is included on the accompanying CD-ROM.

Adaptations of this experiment and details of the use of it in assessing practical skills AO3.3 and AO4.4 are given in the Notes on Activities for teachers/technicians.

Types of mixture

Our world is very complex, owing to the vast range of pure substances available and to the variety of ways in which these pure substances can mix with each other. In everyday life, we do not ‘handle’ pure substances very often. The air we breathe is not a single, pure substance – and we could not live in it if it were! Water would be rather tasteless if we drank it pure (distilled).

Each **mixture** must be made from at least two parts, which may be solid, liquid or gas. There are a number of different ways in which the three states can be combined. In some, the states are completely mixed to become one single state or phase – ‘you cannot see the join’. Technically, the term **solution** is used for this type of mixture.

Solid salt dissolves in liquid water to produce a liquid mixture – a salt solution (Figure 2.10). In general terms, the solid that dissolves in the liquid is called the **solute**. The liquid in which the solid dissolves is called the **solvent**. In other types of mixture, the states remain separate. One phase is

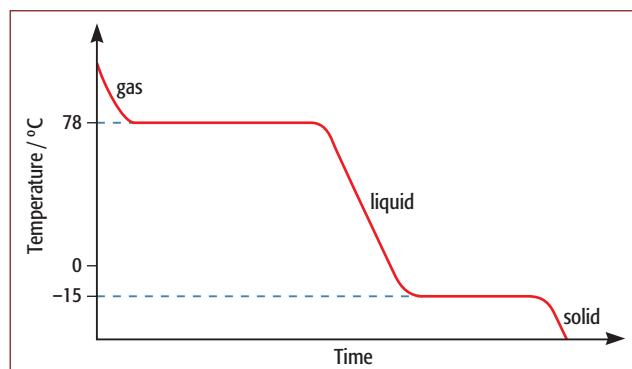


Figure 2.9 The cooling curve for a substance. The temperature stays constant while the gas condenses, and while the liquid freezes. A cooling mixture of ice and salt could be used to lower the temperature below 0 °C.

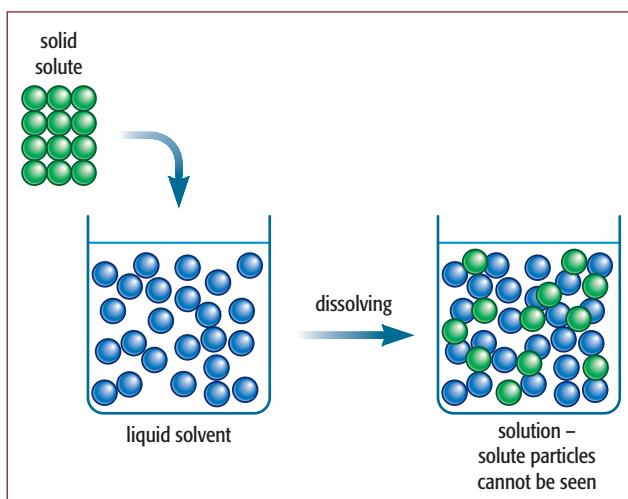


Figure 2.10 When a solute dissolves in a solvent, the solute particles are completely dispersed in the liquid.

broken up into small particles, droplets, or bubbles, within the main phase. Perhaps the most obvious example of this type of mixture is a **suspension** of fine particles of a solid in a liquid, such as we often get after a precipitation reaction.

Solutions

There are various ways in which substances in different states can combine. Perhaps the most important idea here is that of one substance dissolving in another – the idea of a **solution**. We most often think of a solution as being made of a solid dissolved in a liquid. Two-thirds of the Earth's surface is covered by a solution of various salts in water. The salts are totally dispersed in the water and cannot be seen. However, other substances that are not normally solid are dissolved in seawater. For example, the dissolved gases, oxygen and carbon dioxide, are important for life to exist in the oceans.

Less obvious perhaps, but quite common, are solutions of one liquid in another. Alcohol mixes (dissolves) completely with water. Beer, wine and whisky do not separate out into layers of alcohol and water (even when the alcohol content is quite high). Alcohol and water are completely **miscible**: they make a solution.

Alloys are similar mixtures of metals, though we do not usually call them solutions. They are made by mixing the liquid metals together (dissolving one metal in the other) before solidifying the alloy.

Questions

- 2.1 Give the names for the following physical changes:
 - a liquid to solid
 - b liquid to gas at a precise temperature
 - c gas to liquid
 - d solid to gas directly.
- 2.2 What effect does the presence of an impurity have on the freezing point of a liquid?
- 2.3 Sketch a cooling curve for water from 80 °C to –20 °C, noting what is taking place in the different regions of the graph.
- 2.4 What do you understand by the word **volatile** when used in chemistry?
- 2.5 Put these three liquids in order of **volatility**, with the most volatile first: water (b.p. 100 °C), ethanoic acid (b.p. 128 °C), ethanol (b.p. 78 °C).

2.2 Separating and purifying substances

To make sense of the material world around us, we need methods for physically separating the many and varied mixtures that we come across. Being able to purify and identify the many substances present in these mixtures not only satisfies our curiosity but is crucial to our well-being and health. There is a range of physical techniques available to make the necessary separations (Table 2.3, overleaf). They all depend in some way on a difference in the physical properties of the substances in the mixture.

The most useful separation method for a particular mixture depends on:

- ◆ the type of mixture, and
- ◆ which substance in the mixture we are most interested in.

Separating insoluble solids from liquids

In some ways these are the easiest mixtures to separate. Quite often, just leaving a suspension of a solid in a liquid to stand achieves a separation – especially if the particles of solid are large enough. Once the solid has

Mixture	Method of separation
solid + solid (powdered mixture)	use some difference in properties, e.g. density, solubility, sublimation, magnetism
suspension of solid in liquid	filtration or centrifugation
liquid + liquid (immiscible)	use a separating funnel or decantation
solution of solid in liquid	to obtain solid: use evaporation (crystallisation) to obtain liquid: use distillation
two (or more) liquids mixed together (miscible)	fractional distillation
solution of two (or more) solids in a liquid	chromatography

Table 2.3 Separating different types of mixture.

settled to the bottom, the liquid can be carefully poured off – a process called **decanting**.

A more generally useful method for separating solids from liquids is **filtration** (Figure 2.11a). Here the insoluble material is collected as a **residue** on filter paper. Filtration is useful because both phases can be obtained in one process. The liquid phase is collected as the **filtrate**. The process can be speeded up by using a vacuum pump to ‘suck’ the liquid through the filter paper in a Buchner funnel and flask (Figure 2.11b).

Various large-scale filtration methods are used in industry. Perhaps the most useful of these are the filter beds used to treat water for household use.

Another method of separating an insoluble solid from a liquid is **centrifugation** where the mixture is spun at high speed in a centrifuge. This causes the solid to be deposited at the bottom of the centrifuge tube. The liquid can be carefully decanted off.

Separating immiscible liquids

Mixtures of two **immiscible** liquids can be separated if the mixture is placed in a separating funnel and allowed to stand. The liquids separate into different layers. The lower, denser layer is then ‘tapped’ off at the bottom (Figure 2.12). This type of separation is

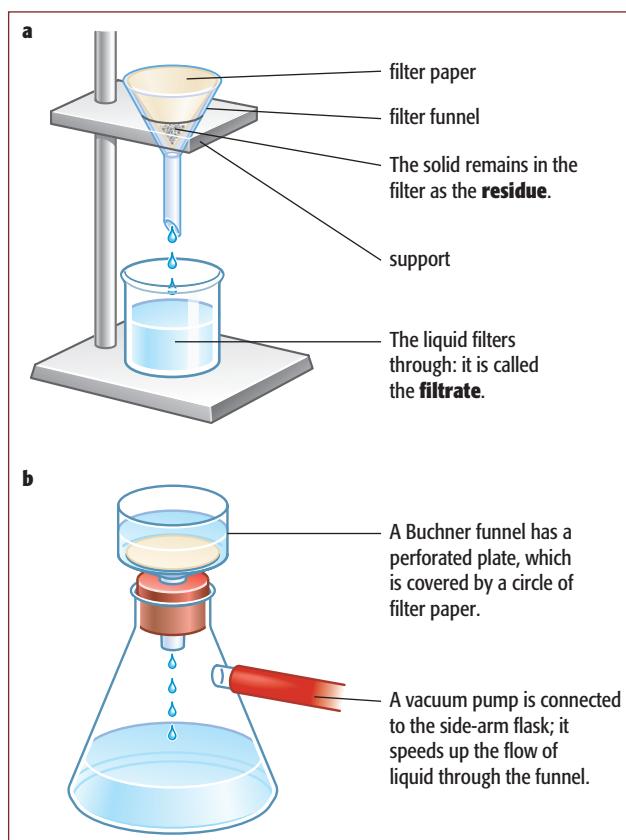


Figure 2.11 Filtration separates an insoluble solid from a liquid.

useful in industry. For example, at the base of the blast furnace the molten slag forms a separate layer on top of the liquid iron. The two can then be ‘tapped’ off separately.

Separating mixtures of solids

The separation of a solid from a mixture of solids depends largely on the particular substance being purified. Some suitable difference in physical properties needs to be found. Usually it helps if the mixture is ground to a powder before any separation is attempted.

Separations based on differences in density

‘Panning’ for gold is still carried out in the search for new deposits. In Amazonia, river-beds are mechanically sifted (‘vacuum-cleaned’) to collect gold dust. These methods depend on the gold dust being denser than the other substances in the river sediment. This type of method is also used in purifying the ores of zinc and copper, although in these cases the metals are less dense than the ores and so float on the surface.

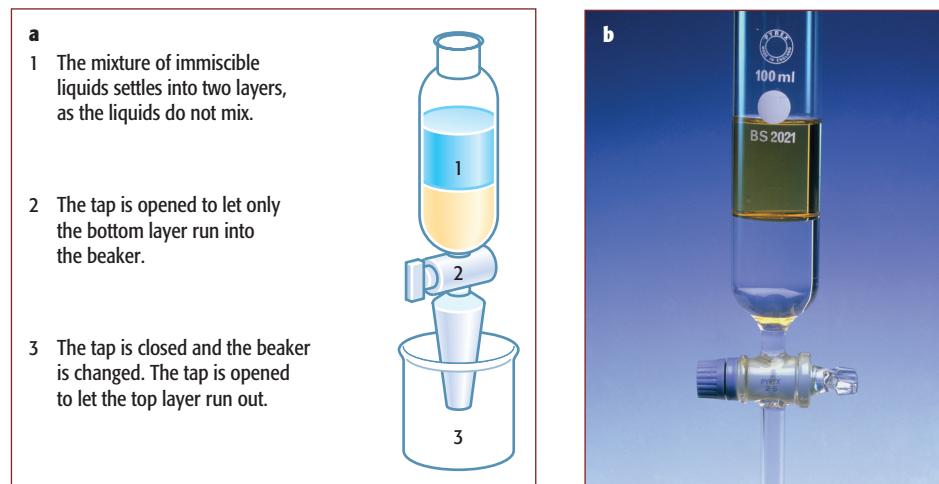


Figure 2.12 **a** A separating funnel can be used to separate two immiscible liquids. **b** Oil and water in a separating funnel.

Separtions based on magnetic properties

Magnetic iron ore can be separated from other material in the crushed ore by using an electromagnet. In the Amazonian gold diggings, magnets are used to clean away iron-containing, red-brown dust from the powdered gold. In the environmentally and economically important processes of recycling metals, iron objects can be picked out from other scrap metal using electromagnets.

Activity 2.2

Separating common salt and sand

Skills

A03.1 Using techniques, apparatus and materials

The aim of this activity is to separate a mixture of salt and sand. The method uses the difference in solubility of the two solids and the technique of filtration.

A worksheet, with a self-assessment checklist, is included on the accompanying CD-ROM.

Separtions based on differences in solubility

One very useful way of separating a soluble substance from a solid mixture is as follows. The mixture is first ground to a powder. A suitable liquid solvent is added. The solvent must dissolve one of the solid substances present, but not the others. The solvent is often water,

but other liquids can be useful. The mixture in the solvent is then warmed and stirred. Care must be taken at the warming stage when using solvents other than water. The warm mixture is then filtered (Figure 2.11). This leaves the insoluble substances as a residue on the filter paper, which can be dried. The soluble substance is in the liquid filtrate. Dry crystals can be obtained by evaporation and **crystallisation**, see Figure 2.14.

Separtions based on sublimation

A solid that sublimes can be separated from others using this property (Figure 2.13).

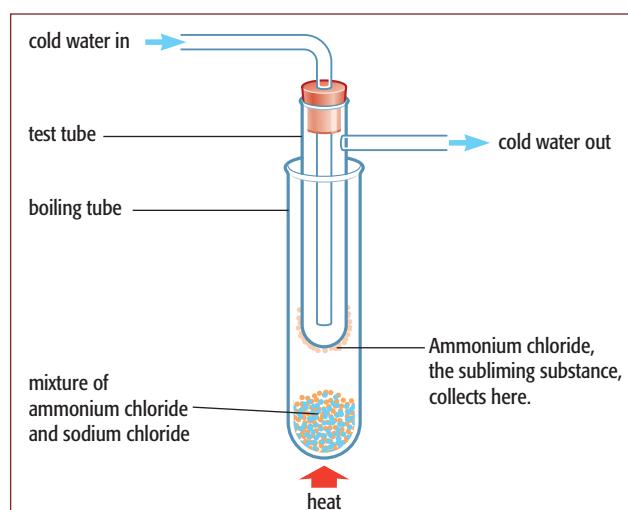


Figure 2.13 Ammonium chloride can be separated from a mixture because it sublimes. The crystals condense on the cooled surface.

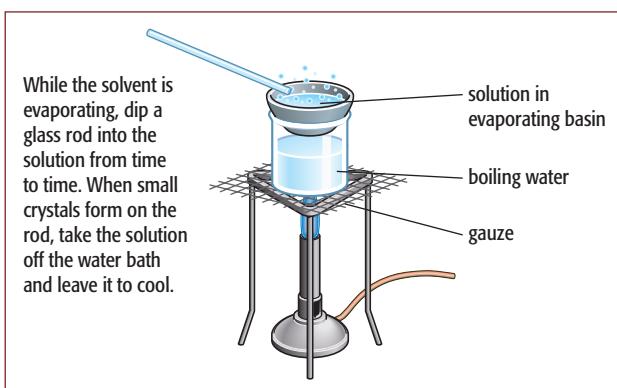


Figure 2.14 An evaporation method. This method should not be used if the solvent is flammable. Instead, use an electrical heating element and an oil or water bath.

Separating solutions

The separation of this type of mixture is often slightly more complicated because there is no physical separation of the phases in the original mixture. The methods of separation usually depend on solubility properties or on differences in boiling point (or volatility).

Separating a solid from solution in a liquid can be carried out by evaporation or crystallisation. Evaporation gives only a powder, but crystallisation can result in proper crystals. Both processes begin by evaporating away the liquid but, when crystals are needed, evaporation is stopped when the solution has been concentrated enough. Figure 2.14 shows how this can be judged and done safely. The concentrated solution is allowed to cool slowly. The crystals formed can then be filtered off and dried.

Separating a liquid from a solution is usually carried out by **distillation** (Figure 2.15). The boiling point of

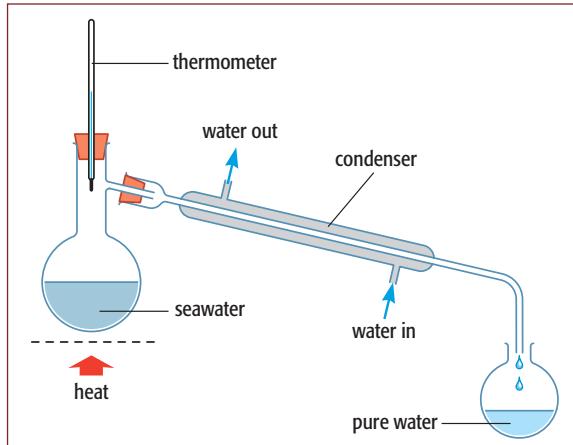


Figure 2.15 The distillation of seawater.

the liquid is usually very much lower than that of the dissolved solid. The liquid is more volatile than the dissolved solid and can easily be evaporated off in a distillation flask. It is condensed by passing it down a water-cooled condenser, and then collected as the **distillate**.

Separating the liquids from a mixture of two (or more) miscible liquids is again based on the fact that the liquids will have different boiling points. However, the boiling points are closer together than for a solid-in-liquid solution and **fractional distillation** must be used (Figure 2.16). In fractional distillation the most volatile liquid in the mixture distils over first and the least volatile liquid boils over last.

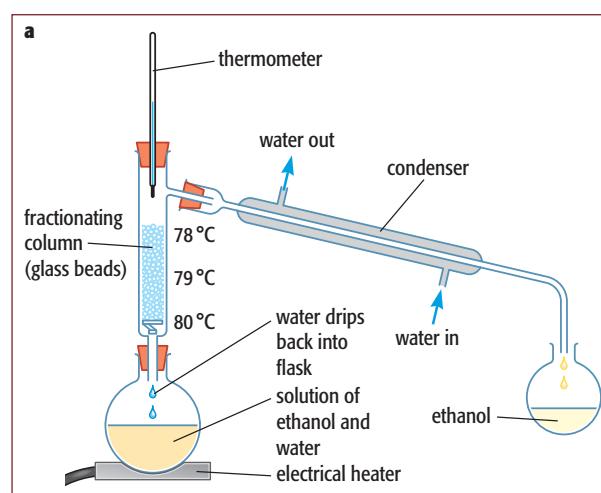


Figure 2.16 **a** Separating a mixture of ethanol (alcohol) and water by fractional distillation. **b** A close-up of the glass beads in the column.

For example, ethanol boils at 78 °C whereas water boils at 100 °C. When a mixture of the two is heated, ethanol and water vapours enter the **fractionating column**. Glass beads in the column provide a large surface area for condensation (Figure 2.16b). Evaporation and condensation take place many times as the vapours rise up the column. Ethanol passes through the condenser first as the temperature of the column is raised above its boiling point. Water condenses in the column and flows back into the flask because the temperature of the column is below its boiling point of 100 °C.

The temperature on the thermometer stays at 78 °C until all the ethanol has distilled over. Only then does the temperature on the thermometer rise to 100 °C and the water distil over. By watching the temperature carefully, the two liquids (**fractions**) can be collected separately.

Fractional distillation is used to separate any solution containing liquids with different boiling points. The liquid in the mixture with the lowest boiling point (the most volatile) distils over first. The final liquid to distil over is the one with the highest boiling point (the least volatile). Fractional distillation can be adapted as a continuous process and is used industrially to separate:

- ◆ the various fractions from petroleum (page 276),
- ◆ the different gases from liquid air (page 7).

Study tip

In fractional distillation remember that it is the liquid with the lowest boiling point that distils over first.

Activity 2.3 Distillation of mixtures

In this experiment, several mixtures will be separated using different types of distillation apparatus, including a microscale distillation apparatus.

A worksheet is included on the CD-ROM.

Separating two or more dissolved solids in solution can be carried out by **chromatography**. There are several types of chromatography, but they all follow the same basic principles. Paper chromatography is probably the simplest form to set up and is very useful if we want to analyse the substances present in a solution. For example, it can tell us whether a solution has become contaminated. This can be very important because contamination of food or drinking water, for instance, may be dangerous to our health.

A drop of concentrated solution is usually placed on a pencil line near the bottom edge of a strip of chromatography paper. The paper is then dipped in the solvent. The level of the solvent must start below the sample. Figure 2.17 (overleaf) shows the process in action.

Many different solvents are used in chromatography. Water and organic solvents (carbon-containing solvents) such as ethanol, ethanoic acid solution and propanone are common. Organic solvents are useful because they dissolve many substances that are insoluble in water. When an organic solvent is used, the process is carried out in a tank with a lid to stop the solvent evaporating.

Activity 2.4 Investigation of food dyes by chromatography

This experiment involves testing some food colours with paper chromatography to find out if they are pure colours or mixtures of several dyes. These food colours are used in cake making, for instance, and there is quite a wide range of permitted colours readily available.

A worksheet is included on the CD-ROM.

Adaptations of this experiment are given in the Notes on Activities for teachers/technicians.

The substances separate according to their solubility in the solvent. As the solvent moves up the paper, the substances are carried with it and begin to separate. The substance that is most soluble moves fastest up the paper. An insoluble substance would remain at the origin. The run is stopped just before the **solvent front** reaches the top of the paper.

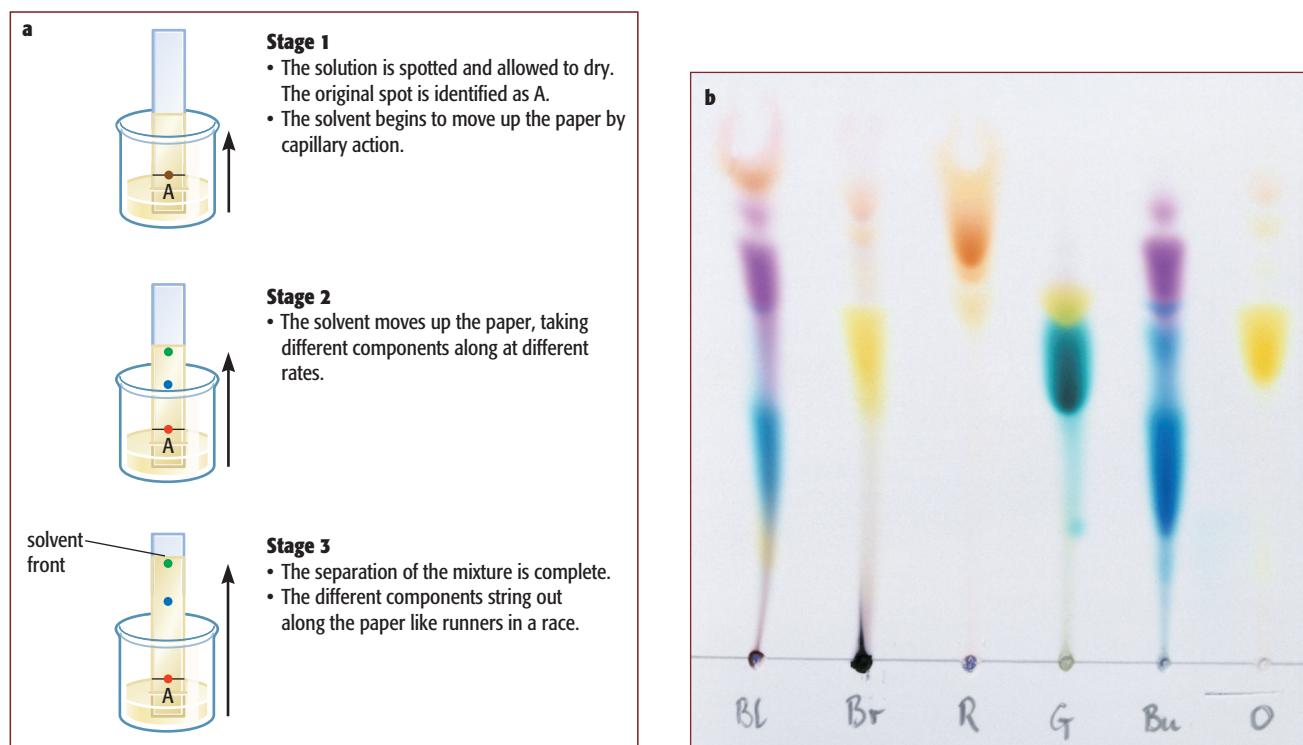


Figure 2.17 **a** Various stages during paper chromatography. The sample is separated as it moves up the paper. **b** A paper chromatogram.

S The distance moved by a particular spot is measured and related to the position of the solvent front. The ratio of these distances is called the ***R_f* value**, or retention factor. This value is used to identify the substance:

$$R_f = \frac{\text{distance moved by the substance}}{\text{distance moved by the solvent front}}$$

Originally, paper chromatography was used to separate solutions of coloured substances (dyes and pigments) since they could be seen as they moved up the paper. However, the usefulness of chromatography has been greatly increased by the use of **locating agents** (Figure 2.18). These mean that the method can also be used for separating substances that are **not** coloured. The paper is treated with locating agent after the chromatography run. The agent reacts with the samples to produce coloured spots.

Chromatography has proved very useful in the analysis of biologically important molecules such as sugars, amino acids and nucleotide bases. In fact, molecules such as amino acids can be ‘seen’ if the paper **chromatogram** is viewed under ultraviolet light.

The purity and identity of substances

Paper chromatography is one test that can be used to check for the **purity** of a substance. If the sample is pure, it should only give **one** spot when run in several different solvents. The **identity** of the sample can also be checked by comparing its *R_f* value to that of a sample we know to be pure.

Probably the most generally used tests for purity are measurements of melting point or boiling point. As we saw earlier, impurities would lower the melting point or raise the boiling point of the substance. They would also make these temperatures less precise. These temperatures have been measured for a very wide range of substances. The identity of an unknown substance can be found by checking against these measured values for known pure substances.

The process of purification is of crucial importance in many areas of the chemical industry. Medicinal **drugs** (pharmaceuticals) must be of the highest possible purity. Any contaminating substances, even in very small amounts, may have harmful side effects. Coloured dyes (food colourings) are added to food and drinks to improve their appearance. The colourings added need to be carefully controlled. In Europe the

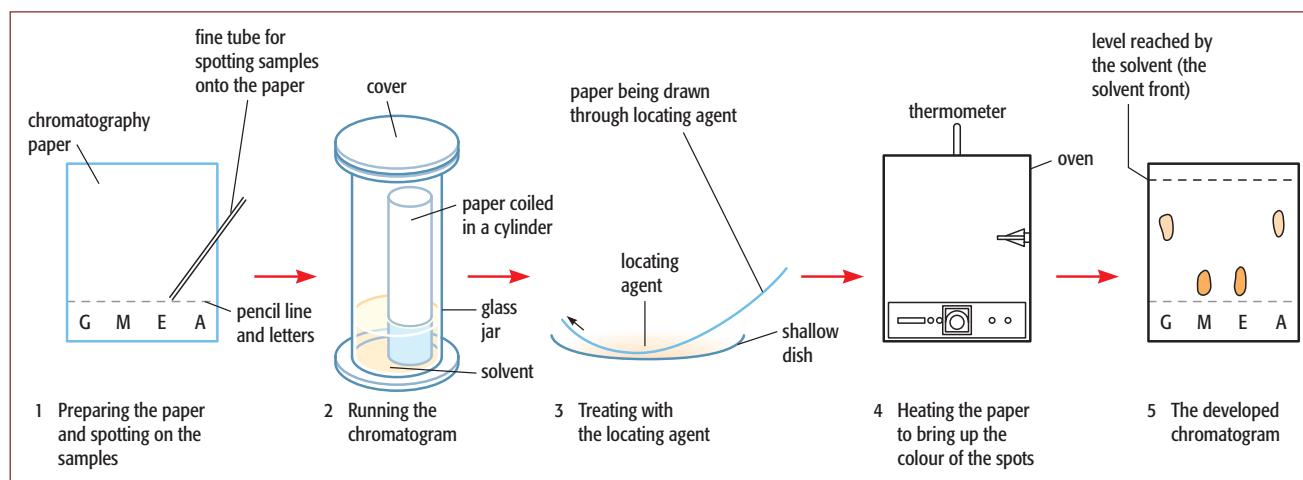


Figure 2.18 Chromatography using a locating agent to detect the spots on the paper. Alternatively, the locating agent can be sprayed on the paper.

permitted colourings are listed as E100 to E180. Many dyes that were once added are now banned. Even those which are permitted may still cause problems for some people. The yellow colouring tartrazine (E102) is found in many drinks, sauces, sweets and snacks. To most people it is harmless, but in some children it appears to cause hyperactivity and allergic reactions, for example asthma. Even where there is overall government regulation, individuals need to be aware of how particular foods affect them.

A closer look at solutions

The solubility of solids in liquids

Probably the most important and common examples of mixtures are solutions of solids in liquids.

Key definition

A **solution** is made up of two parts:

- ◆ the **solute** – the solid that dissolves
- ◆ the **solvent** – the liquid in which it dissolves.

Water is the commonest solvent in use, but other liquids are also important. Most of these other solvents are organic liquids, such as ethanol, propanone and trichloroethane. These organic solvents are important because they will often dissolve substances that do not dissolve in water. If a substance dissolves in a solvent, it is said to be **soluble**: if it does not dissolve, it is **insoluble**.

If we try to dissolve a substance such as copper(II) sulfate in a fixed volume of water, the solution

becomes more concentrated as we add more solid. A **concentrated** solution contains a high proportion of solute; a **dilute** solution contains a small proportion of solute. The **concentration** of a solution is the mass of solute dissolved in a particular volume of solvent, usually 1 dm³.

If we keep adding more solid, a point is reached when no more will dissolve at that temperature. This is a **saturated solution**. To get more solid to dissolve, the temperature must be increased. The concentration of solute in a saturated solution is the **solubility** of the solute at that temperature.

The solubility of most solids increases with temperature. The process of crystallisation depends on these observations. When a saturated solution is cooled, it can hold less solute at the lower temperature, and some solute crystallises out.

The solubility of gases in liquids

Unlike most solids, gases become less soluble in water as the temperature rises. The solubility of gases from the air in water is quite small, but the amount of dissolved oxygen is enough to support fish and other aquatic life. Interestingly, oxygen is more soluble in water than nitrogen is. So when air is dissolved in water, the proportions of the two gases become 61% nitrogen and 37% oxygen. This is an enrichment in life-supporting oxygen compared to air (78% nitrogen and 21% oxygen).

The solubility of gases increases with pressure. Sparkling drinks contain carbon dioxide dissolved under pressure. They ‘fizz’ when the pressure is released

by opening the container. They go ‘flat’ if the container is left to stand open, and more quickly if left to stand in a warm place.

Carbon dioxide is more soluble than either nitrogen or oxygen. This is because it reacts with water to produce carbonic acid. The world is not chemically static. Substances are not only **mixing** with each other but also chemically **reacting**. This produces a world that is continuously changing. To gain a better understanding of this, we need to look more deeply into the ‘make-up’ of chemical substances.

Questions

- 2.6** How would you separate the following?
- water from seawater
 - ethanol from an ethanol/water mixture
 - sugar crystals from a sugar solution
- 2.7** What do you understand by the term **sublimation**?
- 2.8** What type of substance was chromatography originally designed to separate?
- 2.9** How can we now extend the use of chromatography to separate colourless substances?
- 2.10** Define the term R_f value in connection with chromatography.

2.3 Atoms and molecules

Elements and compounds

Earlier in this chapter you were introduced to pure substances, and to ways of purifying and identifying them. But what are ‘pure substances’?

Key definition

There are two types of **pure substance** – elements and compounds:

- ◆ **elements** – substances that cannot be chemically broken down into simpler substances
- ◆ **compounds** – pure substances made from two, or more, elements chemically combined together.

Figure 2.19 summarises what we now know about matter in simple terms. Elements are the ‘building blocks’ from which the Universe is constructed. There are over 100 known elements, but most of the Universe consists of just two. Hydrogen (92%) and helium (7%) make up most of the mass of the Universe, with all the other elements contributing only 1% to the total. The concentration, or ‘coming together’, of certain of these elements to make the Earth is of great interest and significance. There are a total of 94 elements found naturally on Earth but just eight account for more than 98% of the mass of the Earth’s crust. Two elements, silicon and oxygen, which are bound together in silicate rocks, make up almost three-quarters of the crust. Only certain elements are able to form the complex compounds that are found in living things. For example, the human body contains 65% oxygen, 18% carbon, 10% hydrogen, 3% nitrogen, 2% calcium and 2% of other elements.

Chemical reactions and physical changes

Substances can mix in a variety of ways, and they can also react chemically with each other. In a reaction, one substance can be transformed (changed) into another. Copper(II) carbonate is a green solid, but on heating it is changed into a black powder (Figure 2.20). Closer investigation shows that the gas carbon dioxide is also produced. This type of chemical reaction, where a compound breaks down to form two or more substances, is known as **decomposition**.

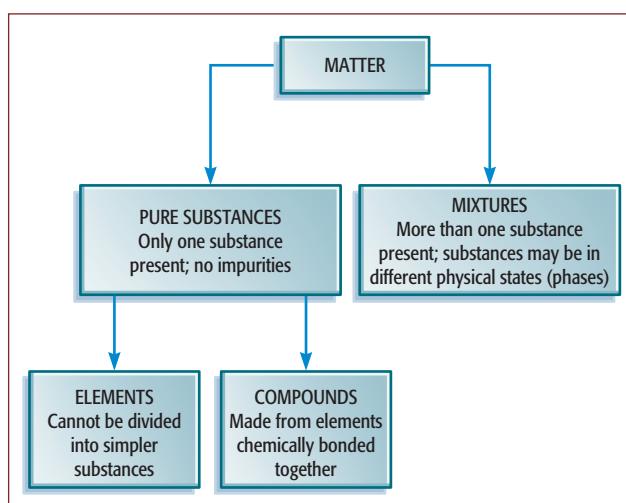


Figure 2.19 Schematic representation of the different types of matter, including elements and compounds.

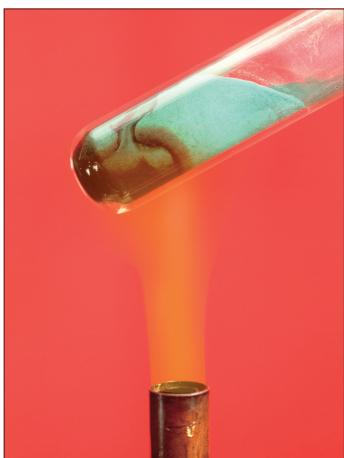


Figure 2.20 Heating copper(II) carbonate.

Decomposition can also be brought about by electricity. Some substances, although they do not conduct electricity when solid, **do** conduct when they are melted or in solution. In the process of conduction, they are broken down into simpler substances. Thus, lead(II) bromide, which is a white powder, can be melted. When a current is passed through molten lead(II) bromide, a silver-grey metal (lead) and a brown vapour (bromine) are formed. Neither of these products can be split into any simpler substances.

The opposite type of reaction, where the substance is formed by the combination of two or more other substances, is known as **synthesis**. For example, if a piece of burning magnesium is plunged into a gas jar of oxygen, the intensity (brightness) of the brilliant white flame increases. When the reaction has burnt out, a white ash remains (Figure 2.21). The ash has totally different properties from the original silver-grey metal strip and colourless gas we started with. A new compound, magnesium oxide, has been formed from magnesium and oxygen.

Although many other reactions are not as spectacular as this, the burning of magnesium shows the general features of chemical reactions.

In a **chemical reaction**:

- ◆ new chemical substance(s) are formed
- ◆ usually the process is not easily reversed
- ◆ energy is often given out.

These characteristics of a chemical reaction contrast with those of a simple physical change such as melting



Figure 2.21 Burning magnesium produces a brilliant white flame.

or dissolving. In a **physical change** the substances involved do not change identity. They can be easily returned to their original form by some physical process such as cooling or evaporation. Sugar dissolves in water, but we can get the solid sugar back by evaporating off the water.

Another synthesis reaction takes place between powdered iron and sulfur. The two solids are finely ground and well mixed. The mixture is heated with a Bunsen burner. The reaction mixture continues to glow after the Bunsen burner is removed. Heat energy is given out. There has been a reaction and we are left with a black non-magnetic solid, iron(II) sulfide, which cannot easily be changed back to iron and sulfur. This example also illustrates some important differences between a mixture (in this case the powders of iron and sulfur) and a compound (in this case the final product of the reaction). The general differences between making a mixture of substances and forming a new compound are shown in Table 2.4, overleaf.

Atomic theory

Elements and compounds mix and react to produce the world around us. They produce massive objects such as the ‘gas giants’ (the planets Jupiter and Saturn), and tiny highly structured crystals of solid sugar.

How do the elements organise themselves to give this variety? How can any one element exist in the three different states of matter simply through a change in temperature?

When a mixture forms ...	When a compound forms ...
the substances are simply mixed together; no reaction takes place	the substances chemically react together to form a new compound
the composition of the mixture can be varied	the composition of the new compound is always the same
the properties of the substances present remain the same	the properties of the new compound are very different from those of the elements in it
the substances in the mixture can be separated by physical methods such as filtration, distillation or magnetic attraction	the compound cannot easily be separated into its elements

Table 2.4 The differences between mixtures and pure compounds.

Our modern understanding is based on the atomic theory put forward by John Dalton in 1807. His theory reintroduced the ideas of Democritus (460–370 BCE) and other Greek philosophers who suggested that all matter was infinitely divided into very small particles known as **atoms**. These ideas were not widely accepted at the time. They were only revived when Dalton developed them further and experimental observations under the microscope showed the random motion of dust particles in suspension in water or smoke particles in air (Brownian motion).

Dalton suggested that:

- ◆ a pure element is composed of atoms
- ◆ the atoms of each element are different in size and mass
- ◆ atoms are the smallest particles that take part in a chemical reaction
- ◆ atoms of different elements can combine to make **molecules** of a compound.

Certain parts of the theory may have needed to change as a result of what we have discovered since Dalton's time. However, Dalton's theory was one of the great leaps of understanding in chemistry. It meant that we

could explain many natural processes. Whereas Dalton only had theories for the existence of atoms, modern techniques (such as scanning tunnelling microscopy) can now directly reveal the presence of individual atoms. It has even been possible to create an 'atomic logo' (Figure 2.22) by using individual atoms, and it may soon be possible to 'see' a reaction between individual atoms.

A chemical language

Dalton suggested that each element should have its own **symbol** – a form of chemical shorthand. He could then write the formulae of compounds without writing out the name every time. Our modern system uses letters taken from the name of the element. This is an international code. Some elements have been known for a long time and their symbol is taken from their Latin name.

The symbol for an element consists of one or two letters. Where the names of several elements begin with the same letter, the second letter of the name is usually included in lower case (Table 2.5). As more elements were discovered, they were named after a wider range of people, cities, countries and even particular universities. We shall see in Chapter 3 how useful it is to be able to use symbols, and how they can be combined to show the formulae of complex chemical compounds. A full list of the elements and their symbols is found in the Periodic Table (see the Appendix).

The kinetic theory of matter

The idea that all substances consist of very small particles begins to explain the structure of the three different states of matter. The kinetic theory of matter describes these states, and the changes between them, in terms of the movement of particles.

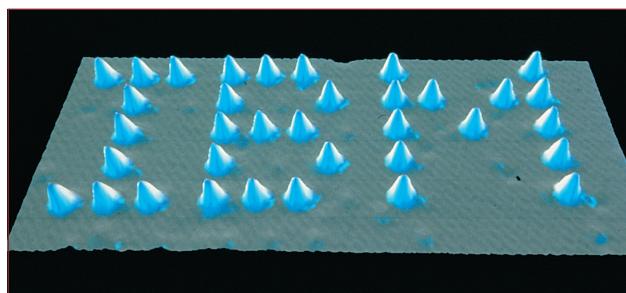


Figure 2.22 An 'atomic logo' produced by xenon atoms on a nickel surface 'seen' using scanning tunnelling microscopy.

Element	Latin name	Symbol
hydrogen		H
helium		He
carbon		C
calcium		Ca
copper	cuprum	Cu
chlorine		Cl
nitrogen		N
sodium	natrium	Na
phosphorus		P
potassium	kalium	K
iron	ferrum	Fe
lead	plumbum	Pb
silver	argentum	Ag
gold	aurum	Au

Table 2.5 The symbols of some chemical elements.

The main points of the kinetic theory

- ◆ All matter is made up of very small particles (different substances contain different types of particles – such as atoms or molecules).
- ◆ The particles are moving all the time (the higher the temperature, the higher the average energy of the particles). In a gas, the faster the particles are moving, the higher the temperature.
- ◆ The freedom of movement and the arrangement of the particles is different for the three states of matter (Figure 2.23).
- ◆ The pressure of a gas is produced by the atoms or molecules of the gas hitting the walls of the container. The more often the particles collide with the walls, the greater the pressure.

Figure 2.23 is a summary of the organisation of the particles in the three states of matter, and helps to explain their different overall physical properties. The highly structured, ordered microscopic arrangements (**lattices**) in solids can produce the regular crystal structures seen in this state. The ability of the particles to move in the liquid and gas phases

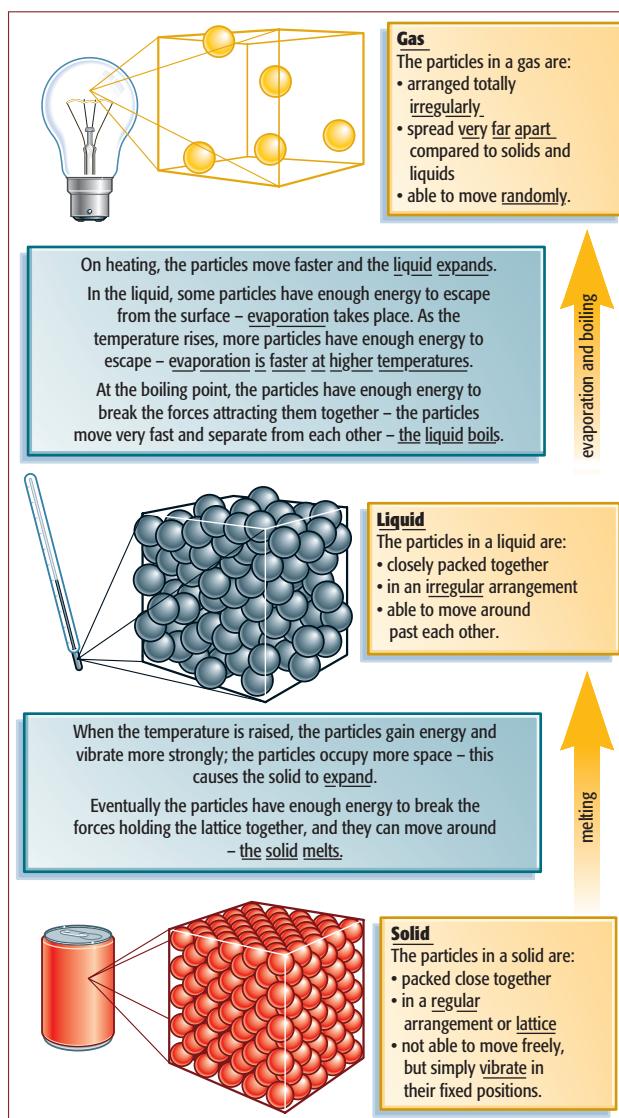


Figure 2.23 Applying the kinetic theory to changes in physical state.

produces their fluid properties. The particles are very widely separated in a gas, but are close together in a liquid or solid. The space between the particles can be called the **intermolecular space** (IMS). In a gas, the intermolecular space is large and can be reduced by increasing the external pressure – gases are compressible. In liquids, this space is very much smaller – liquids are not very compressible.

Study tip

It's important to realise that even in a liquid, the particles are still close together, although they can move around and past each other.

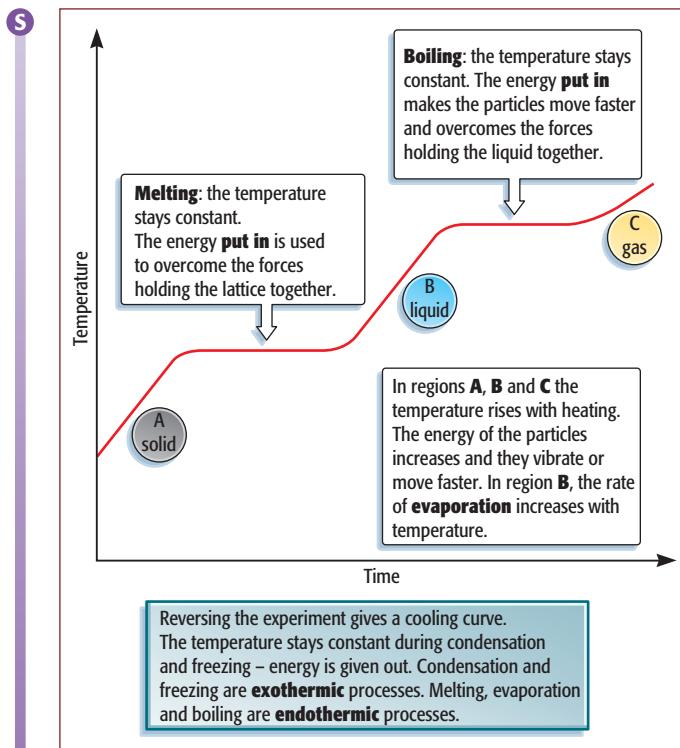


Figure 2.24 The energy changes taking place during heating and cooling.

The way the particles in the three states are arranged also helps to explain the temperature changes when a substance is heated or cooled. Figure 2.24 (overleaf) summarises the energy changes taking place at the different stages of a heating-curve or cooling-curve experiment.

Diffusion in fluids

The idea that fluids are made up of moving particles helps us to explain processes involving **diffusion**.

Dissolving

A potassium manganate(vii) crystal is placed at the bottom of a dish of water. It is then left to stand. At first the water around the crystal becomes purple as the solid dissolves (Figure 2.25). Particles move off the surface of the crystal into the water. Eventually the crystal dissolves completely and the whole solution becomes purple. The particles from the solid become evenly spread through the water.

Whether a solid begins to break up like this in a liquid depends on the particular solid and liquid involved.

But the spreading of the solute particles throughout the liquid is an example of diffusion. Diffusion in solution is also important when the solute is a gas. This is especially important in breathing! Diffusion contributes to the



Figure 2.25 The diffusion of potassium manganate(vii) in water as it dissolves.

movement of oxygen from the lungs to the blood, and of carbon dioxide from the blood to the lungs.

The diffusion of gases

A few drops of liquid bromine are put into a gas jar and the lid is replaced. After a short time the jar becomes full of brown gas. Bromine vaporises easily and its gas will completely fill the container (Figure 2.26). Gases diffuse to fill all the space available to them. Diffusion is important for our ‘sensing’ of the world around us. It is the way smells reach us, whether they are pleasant or harmful.

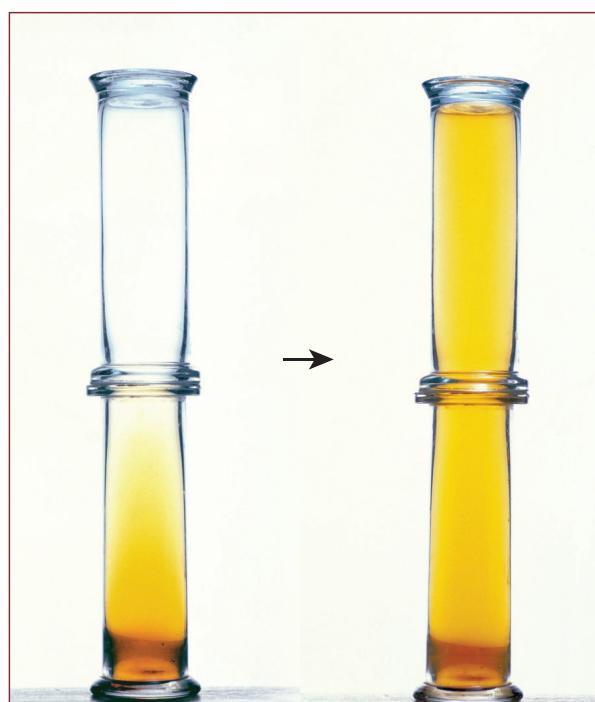


Figure 2.26 Bromine vapour diffuses (spreads) throughout the container to fill all the space.

Study tip

The key idea about diffusion is the idea of particles spreading to fill the space available to the molecules.

Key definition

diffusion – the process by which different fluids mix as a result of the random motions of their particles.

- ◆ Diffusion involves the movement of particles from a region of higher concentration towards a region of lower concentration. Eventually, the particles are evenly spread – their concentration is the same throughout.
- ◆ It does not take place in solids.
- ◆ Diffusion in liquids is much slower than in gases.

S Not all gases diffuse at the same rate. This is shown by the experiment in Figure 2.27. The ammonia and hydrochloric acid fumes react when they meet,

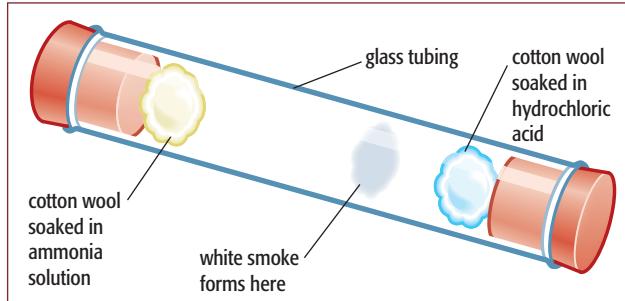


Figure 2.27 Ammonia and hydrochloric acid fumes diffuse at different rates.

producing a white ‘smoke ring’ of ammonium chloride. The fact that the ring is not formed halfway along the tube shows that ammonia, the lighter molecule of the two, diffuses faster.

The speeds of gas atoms or molecules are high. We are being bombarded constantly by nitrogen and oxygen molecules in the air, which are travelling at about 500 m/s (1800 km/h). However, these particles collide very frequently with other particles in the air (many millions of collisions per second), so their path is not particularly direct! (Figure 2.28a). These very frequent collisions slow down the overall rate of diffusion from one place to another.

The movement of individual gas molecules or atoms in the air cannot be seen: the particles are far too small.

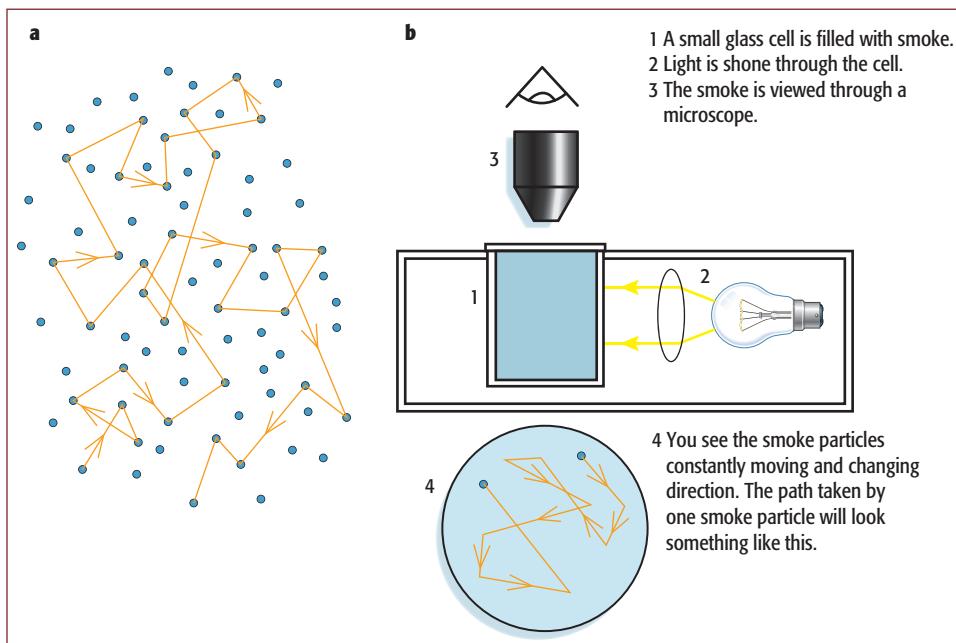


Figure 2.28 a Diffusion of an individual gas molecule or atom; the particle collides with many others, deflecting its path. b Demonstrating Brownian motion using a smoke cell; the smoke particles show a random motion.

S However, the effect of their presence and motion can be shown using a smoke cell (Figure 2.28b). The smoke particles are hit by the invisible molecules in the air. The jerky, random motion produced by these hits can be seen under a microscope. This is known as **Brownian motion**.

Three important points derived from **kinetic theory** are relevant here:

- ◆ Heavier particles move more slowly than lighter particles at the same temperature; larger molecules diffuse more slowly than smaller ones.
- ◆ The pressure of a gas is the result of collisions of the fast-moving particles with the walls of the container.
- ◆ The average speed of the particles increases with an increase in temperature.

Activity 2.5 Investigating diffusion – a demonstration

Skills

- A03.1 Using techniques, apparatus and materials
A03.3 Observing, measuring and recording
A03.4 Interpreting and evaluating observations and data

This is the classic demonstration of the diffusion of gases in which ammonia and hydrogen chloride meet in a long tube. The demonstration shows how the progress of the gases can be tracked using indicator. Measurements can be made to give an estimate of the rate of diffusion of the two gases.

Worksheets are included on the accompanying CD-ROM for both the teacher demonstration and a microscale version of the experiment which could be carried out by students.

Details of other demonstrations and experiments on diffusion are given in the Notes on Activities for teachers/technicians.

Atoms and molecules

The behaviour of some gaseous elements (their diffusion and pressure) shows that they are made up of molecules, not separate atoms. This is true of hydrogen (H_2), nitrogen (N_2), oxygen (O_2) and others. But, as we discussed on page 36, Dalton had originally introduced the idea of molecules to explain the particles making up **compounds** such as water, carbon dioxide and methane. Molecules of these compounds consist of atoms of different **elements** chemically bonded together. Water is made up of two atoms of hydrogen bonded to one atom of oxygen, giving the formula H_2O . Methane (CH_4) has one atom of carbon bonded to four atoms of hydrogen, and hydrogen chloride (HCl) has one atom of hydrogen and one atom of chlorine bonded together. Models of these are shown in Figure 2.29.

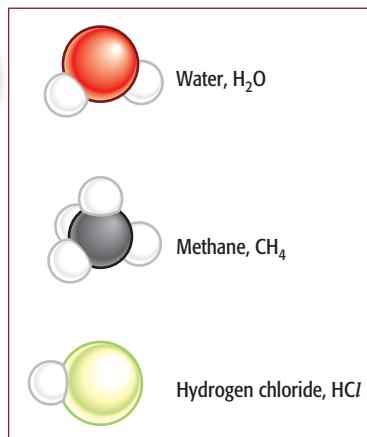


Figure 2.29 Simple compounds consisting of molecules made up of atoms of two different elements.

Questions

- 2.11 Define an element.
- 2.12 Define a compound.
- 2.13 Summarise the differences between the three states of matter in terms of the arrangement of the particles and their movement.
- 2.14 Which gas diffuses faster, ammonia or hydrogen chloride? Briefly describe an experiment that demonstrates this difference.
- 2.15 Which gas will diffuse fastest of all?

2.4 The structure of the atom

Atomic structure

How can atoms join together to make molecules? What makes certain atoms more ready to do this? Why do hydrogen atoms pair up but helium atoms remain single?

To find answers to questions like these, we need first to consider the structure of atoms in general. Dalton thought they were solid, indivisible particles. But research since then has shown that atoms are made up of various subatomic particles. J. J. Thompson discovered the electron (in 1897) and the proton. Crucial experiments were then carried out in Rutherford's laboratory in Manchester in 1909 that showed that the atom is largely empty space. Rutherford calculated that an atom is mostly space occupied by the negatively charged electrons, surrounding a very small, positively charged **nucleus**. The nucleus is at the centre of the atom and contains almost all the mass of the atom. By 1932, when the neutron was discovered, it was clear that atoms consisted of three **subatomic particles – protons, neutrons and electrons**. These particles are universal – all atoms are made from them. The atom remains the smallest particle that shows the characteristics of a particular element.

Measuring the size of atoms

Modern methods such as scanning tunnelling microscopy have allowed us to 'see' individual atoms in a structure. However, atoms are amazingly small! A magnification of 100 million times is necessary to show the stacking pattern of the atoms that make up a gold bar.

A single atom is so small that it cannot be weighed on a balance. However, the mass of one atom can be compared with that of another using a **mass spectrometer**. The element carbon is chosen as the standard. The masses of atoms of all other elements are compared to the mass of a carbon atom. This gives a series of values of **relative atomic mass** for the elements. Carbon is given a relative atomic mass of exactly 12, which can be written as carbon-12. Table 2.6 gives some examples of the values obtained for other

Activity 2.6 Discovering the structure of the atom

Skills

Research skills

ICT skills

The discovery of the nature of the subatomic particles that make up all atoms took place in a relatively short space of time around the beginning of the twentieth century.

Investigate this key period in the history of science using library and internet sources. Devise a PowerPoint or poster presentation on the significant discoveries and the scientists involved.

Key scientists to research are J. J. Thompson, Ernest Rutherford and James Chadwick.

Questions

- A1 What was remarkable about the structure of the atom suggested by the Geiger–Marsden experiments?
- A2 What is it about the nature of the neutron that made it the last of the particles to be discovered?

Element	Atomic symbol	Relative atomic mass
carbon	C	12
hydrogen	H	1
oxygen	O	16
calcium	Ca	40
copper	Cu	64
gold	Au	197

Table 2.6 The relative atomic masses of some elements.

elements. It shows that carbon atoms are 12 times as heavy as hydrogen atoms, which are the lightest atoms of all. Calcium atoms are 40 times as heavy as hydrogen atoms.

Subatomic particles

To get some idea of just how small the nucleus is in comparison to the rest of the atom, here is a simple comparison. If the atom were the size of a football stadium, the nucleus (at the centre-spot) would be the size of a pea!

Protons and neutrons have almost the same mass. Electrons have virtually no mass at all ($\frac{1}{1840}$ of the mass of a proton). The other important feature of these particles is their electric charge. Protons and electrons have equal and opposite charges, while neutrons are electrically neutral (have no charge). The characteristics of these three subatomic particles are listed in Table 2.7.

A single atom is electrically neutral (it has no overall electric charge). This means that in any atom there must be equal numbers of protons and electrons. In this way the total positive charge on the nucleus (due to the protons) is balanced by the total negative charge of the orbiting electrons. The simplest atom of all has one proton in its nucleus. This is the hydrogen atom. It is the only atom that has no neutrons; it consists of one proton and one electron. Atoms of different elements are increasingly complex.

The next simplest atom is that of helium. This has two protons and two neutrons in the nucleus, and two orbiting electrons (Figure 2.30).

The next, lithium, has three protons, four neutrons and three electrons. The arrangements in the following atoms get more complicated with the addition of more protons and electrons. The number of neutrons required to hold the nucleus together increases as the atomic size increases. Thus, an atom of gold consists of 79 protons (p^+), 118 neutrons (n^0) and 79 electrons (e^-).

Subatomic particle	Relative mass	Relative charge	Location in atom
proton	1	+1	in nucleus
neutron	1	0	in nucleus
electron	$\frac{1}{1840}$ (negligible)	-1	outside nucleus

Table 2.7 Properties of the subatomic particles.

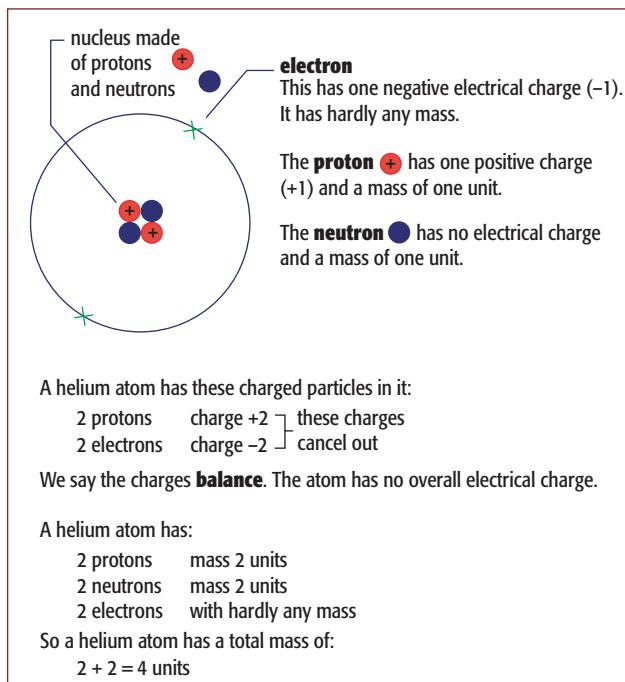


Figure 2.30 The structure of a helium atom.

Proton (atomic) number and nucleon number

Only hydrogen atoms have one proton in their nuclei. Only helium atoms have two protons. Indeed, only gold atoms have 79 protons. This shows that the number of protons in the nucleus of an atom decides which element it is. This very important number is known as the **proton number** (or **atomic number**, given the symbol Z) of an atom.

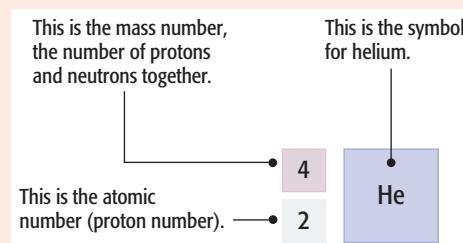
Protons alone do not make up all the mass of an atom. The neutrons in the nucleus also contribute to the total mass. The mass of the electrons can be regarded as so small that it can be ignored. Because a proton and a neutron have the same mass, the mass of a particular atom depends on the total number of protons and neutrons present. This number is called the **nucleon number** (or **mass number**, given the symbol A) of an atom.

The atomic number Z and mass number A of an atom of an element can be written alongside the symbol for that element, in the general way as ${}^A_Z X$. So the symbol for an atom of lithium is ${}^3_3 Li$. The symbols for carbon, oxygen and uranium atoms are ${}^{12}_6 C$, ${}^{16}_8 O$ and ${}^{238}_{92} U$.

If these two important numbers for any atom are known, then its subatomic composition can be worked out.

For proton number and nucleon number we have:

- ◆ proton (atomic) number (Z)
= number of protons in the nucleus
- ◆ nucleon (mass) number (A)
= number of protons + number of neutrons



These two relationships are useful:

- ◆ number of electrons = number of protons
= atomic (proton) number
- ◆ number of neutrons
= nucleon number – proton number
= $A - Z$

Table 2.8 shows the numbers of protons, neutrons and electrons in some different atoms. Note that the rules apply even to the largest, most complicated atom found naturally in substantial amounts.

Study tip

Remember that you can use the Periodic Table you have in the exam for information on these numbers for any atom. Magnesium is the twelfth atom in the table, so it must have 12 protons and 12 electrons in its atoms.

Isotopes

Measurements of the atomic masses of some elements using the mass spectrometer were puzzling. Pure samples of elements such as carbon, chlorine and many others were found to contain atoms with different masses even though they contained the same numbers of protons and electrons. The different masses were caused by different numbers of neutrons in their nuclei. Such atoms are called **isotopes**.

Study tip

Remember that it is just the number of neutrons in the atoms that is the difference between isotopes. They have the same number of protons and electrons.

Atom	Symbol	Atomic number, Z	Mass number, A	Inside the nucleus:		Outside the nucleus: Electrons (Z)
				Protons (Z)	Neutrons ($A - Z$)	
hydrogen	H	1	1	1	0	1
helium	He	2	4	2	2	2
lithium	Li	3	7	3	4	3
beryllium	Be	4	9	4	5	4
carbon	C	6	12	6	6	6
oxygen	O	8	16	8	8	8
sodium	Na	11	23	11	12	11
calcium	Ca	20	40	20	20	20
gold	Au	79	197	79	118	79
uranium	U	92	238	92	146	92

Table 2.8 The subatomic composition and structure of certain atoms.

Element	Isotopes		
Hydrogen	hydrogen (99.99%)	deuterium (0.01%)	tritium ^(a)
	${}_1^1\text{H}$	${}_1^2\text{H}$	${}_1^3\text{H}$
	1 proton	1 proton	1 proton
	0 neutrons	1 neutron	2 neutrons
	1 electron	1 electron	1 electron
Carbon	carbon-12 (98.9%)	carbon-13 (1.1%)	carbon-14 ^(a) (trace)
	${}_6^{12}\text{C}$	${}_6^{13}\text{C}$	${}_6^{14}\text{C}$
	6 protons	6 protons	6 protons
	6 neutrons	7 neutrons	8 neutrons
	6 electrons	6 electrons	6 electrons
Neon	neon-20 (90.5%)	neon-21 (0.3%)	neon-22 (9.2%)
	${}_{10}^{20}\text{Ne}$	${}_{10}^{21}\text{Ne}$	${}_{10}^{22}\text{Ne}$
	10 protons	10 protons	10 protons
	10 neutrons	11 neutrons	12 neutrons
	10 electrons	10 electrons	10 electrons
Chlorine	chlorine-35 (75%)	chlorine-37 (25%)	
	${}_{17}^{35}\text{Cl}$	${}_{17}^{37}\text{Cl}$	
	17 protons	17 protons	
	18 neutrons	20 neutrons	
	17 electrons	17 electrons	

^(a)Tritium and carbon-14 atoms are radioactive isotopes because their nuclei are unstable.

Table 2.9 Several elements that exist as mixtures of isotopes.

Key definition

isotopes – atoms of the same element which have the same proton number but a different nucleon number.

- ◆ The atoms have the same number of protons and electrons, but different numbers of neutrons in their nuclei.
- ◆ Isotopes of an element have the same chemical properties because they have the same electron structure.
- ◆ Some isotopes have unstable nuclei; they are **radioisotopes** and emit various forms of radiation.

The isotopes of an element have the same chemical properties because they contain the same number of electrons. It is the number of electrons in an atom that decides the way in which it forms bonds and reacts with other atoms. However, some physical properties of the isotopes are different. The masses of the atoms differ and therefore other properties, such as density and rate of diffusion, also vary. The modern mass spectrometer shows that most elements have several different isotopes that occur naturally. Others, such as tritium – an isotope of hydrogen (Table 2.9) – can be made artificially.

Tritium and carbon-14 illustrate another difference in physical properties that can occur between isotopes,

as they are **radioactive**. The imbalance of neutrons and protons in their nuclei causes them to be unstable so the nuclei break up spontaneously (that is, without any external energy being supplied), emitting certain types of radiation. They are known as **radioisotopes**.

Relative atomic masses

Most elements exist naturally as a mixture of isotopes. Therefore, the value we use for the atomic mass of an element is an average mass. This takes into account the proportions (abundance) of all the naturally occurring isotopes. If a particular isotope is present in high proportion, it will make a large contribution to the average. This average value for the mass of an atom of an element is known as the relative atomic mass (A_r).

Key definition

relative atomic mass (A_r) – the average mass of normally occurring atoms of an element on a scale where the carbon-12 atom has a mass of exactly 12 units.

Because there are several isotopes of carbon, the standard against which all atomic masses are measured has to be defined precisely. The isotope carbon-12 is used as the standard. One atom of carbon-12 is given the mass of 12 precisely. From this we get that 1 atomic mass unit (a.m.u.) = $\frac{1}{12} \times$ mass of one atom of carbon-12.

The existence of isotopes also explains why most relative atomic masses are not whole numbers. But, to make calculations easier, in this book they are rounded to the nearest whole number. There is one exception, chlorine, where this would be misleading. Chlorine contains two isotopes, chlorine-35 and chlorine-37, in a ratio of 3:1 (or 75%:25%). If the mixture were 50%:50%, then the relative atomic mass of chlorine would be 36. The fact that there is more of the lighter isotope moves the value lower than 36. The actual value is 35.5. The relative atomic mass of chlorine can be calculated by finding the total mass of 100 atoms:

$$\begin{aligned}\text{mass of 100 atoms} &= (35 \times 75) + (37 \times 25) \\ &= 3550\end{aligned}$$

Then,

$$\text{average mass of one atom} = \frac{3550}{100} = 35.5$$

Thus, for chlorine:

$$A_r(\text{Cl}) = 35.5$$

Radioactivity

Some elements have unstable isotopes, such as tritium and carbon-14. The extra neutrons in their nuclei cause them to disintegrate or decay spontaneously. This is **radioactivity** and takes place through nuclear fission. The result of these disintegrations is the release of heat energy and various forms of radioactive radiation. Uranium-235 is a radioactive isotope which is used as a controlled source of energy in nuclear power stations.

The decay is a completely random process and is unaffected by temperature or whether the isotope is part of a compound or present as the free element. Radioactive decay is a nuclear process and not a chemical reaction.

The uses of radioactivity

Radioactive dating

Each radioactive isotope decays at its own rate. However, the time taken for the radioactivity in a sample to halve is constant for a particular radio-isotope. This time is called the **half-life**. Some isotopes have very short half-lives of only seconds: for example, oxygen-14 has a half-life of 71 s. Other half-lives are quite long: for example, carbon-14 has a half-life of 5730 years.

One important use of half-life values is in radioactive dating. Radiocarbon dating (which uses carbon-14) can be used to date wooden and organic objects.

Industrial uses of radioisotopes

Despite the need to handle them with strict safety precautions, radioactive isotopes are widely used in industry and medicine. Most important is the use of an isotope of uranium, ^{235}U , in nuclear power stations. Here, as the isotope splits into smaller parts, as a result

of being bombarded with neutrons, huge amounts of energy are produced. This nuclear fission reaction is the same as that used in the atomic bomb. The difference is that in a nuclear power station, the reaction is controlled.

Other industrial uses of radioisotopes include monitoring the level of filling in containers, checking the thickness of sheets of plastic, paper or metal foil (for example, aluminium baking foil) during continuous production, and detecting leaks in gas or oil pipes (Figure 2.31).

Medical and food-safety uses of radioisotopes

The ease of detection of radioisotopes gives rise to several of their other uses, and in medicine some of their most dangerous properties can be turned to advantage. Several medical uses of radiation depend on the fact that biological cells are sensitive to radioactive emissions. Cells that are growing and dividing are particularly likely to be damaged. Cancer cells are cells that are growing out of control in a tissue of the body. Because of this they are more easily killed by radiation than are healthy

cells. Penetrating γ -radiation from the radioisotope cobalt-60 is used to treat internal cancer tumours. Skin cancer tumours can be treated with less-penetrating radiation. This is done by strapping sheets containing phosphorus-32 or strontium-90 to the affected area of the skin.

Bacterial cells grow and divide rapidly. They are particularly sensitive to radiation. Medical instruments, dressings and syringes can be sterilised by sealing them in polythene bags and exposing them to intense doses of γ -radiation. This has proved a very effective method of killing any bacteria on them.

Study tip

The syllabus specifies that you simply need to know a medical and an industrial use for radioactivity. Do be clear about the difference between radiotherapy and chemotherapy in the treatment of cancer. It is radiotherapy that involves the use of radioactivity to kill the tumour cells.

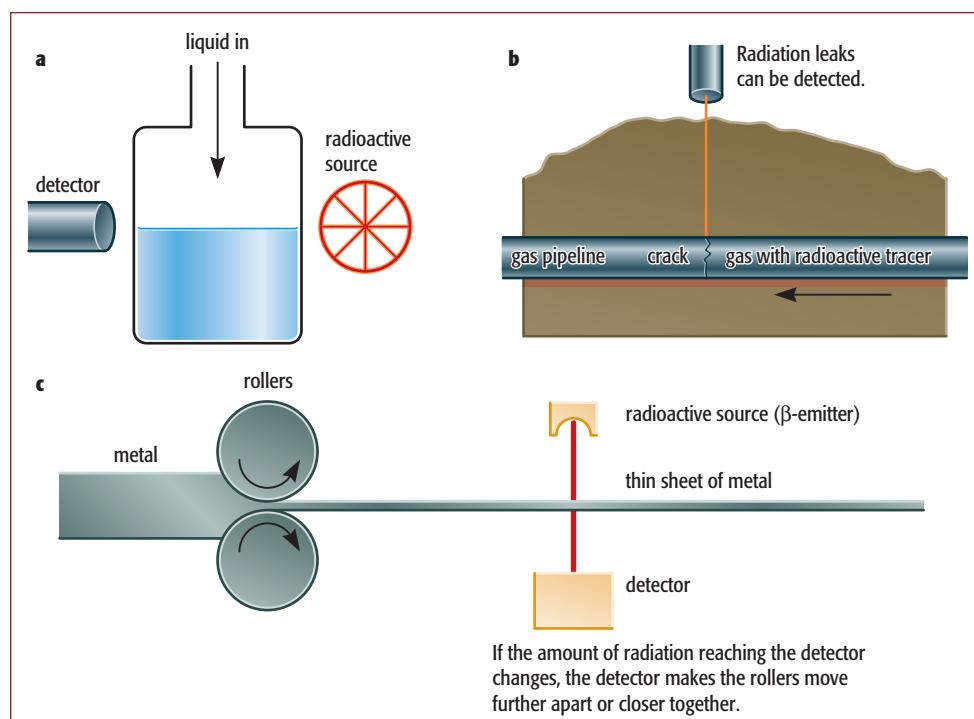


Figure 2.31 Uses of radioactivity: **a** detecting the level of liquid in a container, **b** detecting leaks in underground pipes, and **c** controlling the thickness of metal sheets.



Questions

- 2.16** What are the relative masses of a proton, neutron and electron given that a proton has a mass of 1?
- 2.17** How many protons, neutrons and electrons are there in an atom of phosphorus, which has a proton number of 15 and a nucleon number of 31?
- 2.18** What is the difference in terms of subatomic particles between an atom of chlorine-35 and an atom of chlorine-37?
- 2.19** Give **one** medical and **one** industrial use of radioactivity.

2.5 Electron arrangements in atoms

The aurora borealis (Figure 2.32) is a spectacular display seen in the sky in the far north (a similar phenomenon – the aurora australis – occurs in the night sky of the far south). It is caused by radiation from the Sun moving the electrons in atoms of the gases of the atmosphere.

Similar colour effects can be created in a simpler way in the laboratory by heating the compounds of some metals in a Bunsen flame (see page 208). These colours are also seen in fireworks. The colours produced are due to electrons in the atom moving between two different **energy levels**.

In 1913, Niels Bohr, working with Rutherford in Manchester, developed a theory to explain how



Figure 2.32 The aurora borealis or northern lights as seen from Finland.

electrons were arranged in atoms. This theory helps to explain how the colours referred to above come about.

A simplified version of **Bohr's theory** of the arrangement of electrons in an atom can be summarised as follows (see also Figure 2.33):

- ◆ Electrons are in orbit around the central nucleus of the atom.
- ◆ The electron orbits are called **shells** (or **energy levels**) and have different energies.
- ◆ Shells which are further from the nucleus have higher energies.
- ◆ The shells are filled starting with the one with lowest energy (closest to the nucleus).
- ◆ The first shell can hold only 2 electrons.
- ◆ The second and subsequent shells can hold 8 electrons to give a stable (noble gas) arrangement of electrons.

Other evidence was found that supported these ideas of how the electrons are arranged in atoms. The number and arrangement of the electrons in the atoms of the first 20 elements in the Periodic Table (see the Appendix) are shown in Table 2.10 (overleaf).

When the two essential numbers describing a particular atom are known, the numbers of protons and neutrons, a subatomic picture can be drawn. Figure 2.34 (overleaf) shows such a picture for perhaps the most versatile atom in the Universe, an atom of carbon-12. Studying the organisation of the electrons of an atom is valuable. It begins to explain the patterns in properties of the elements that are the basis of the Periodic Table. This will be discussed in the next chapter.

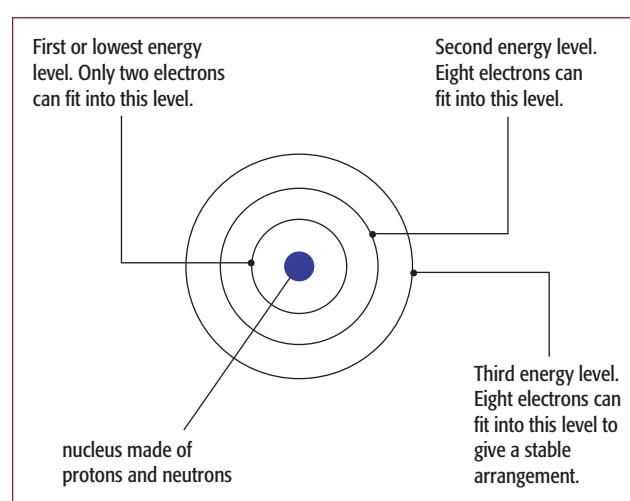


Figure 2.33 Bohr's theory of the arrangement of electrons in an atom.

Element	Symbol	Atomic number, Z	First shell	Second shell	Third shell	Fourth shell	Electron configuration
hydrogen	H	1	●				1
helium	He	2	●●				2
lithium	Li	3	●●	●			2,1
beryllium	Be	4	●●	●●			2,2
boron	B	5	●●	●●●			2,3
carbon	C	6	●●	●●●●			2,4
nitrogen	N	7	●●	●●●●●			2,5
oxygen	O	8	●●	●●●●●●			2,6
fluorine	F	9	●●	●●●●●●●			2,7
neon	Ne	10	●●	●●●●●●●●			2,8
sodium	Na	11	●●	●●●●●●●●●	●		2,8,1
magnesium	Mg	12	●●	●●●●●●●●●●	●●		2,8,2
aluminium	Al	13	●●	●●●●●●●●●●●	●●●		2,8,3
silicon	Si	14	●●	●●●●●●●●●●●●	●●●●		2,8,4
phosphorus	P	15	●●	●●●●●●●●●●●●●	●●●●●		2,8,5
sulfur	S	16	●●	●●●●●●●●●●●●●●	●●●●●●●		2,8,6
chlorine	Cl	17	●●	●●●●●●●●●●●●●●●	●●●●●●●●●		2,8,7
argon	Ar	18	●●	●●●●●●●●●●●●●●●●	●●●●●●●●●●●●●●●		2,8,8
potassium	K	19	●●	●●●●●●●●●●●●●●●●●	●●●●●●●●●●●●●●●●●●	●	2,8,8,1
calcium	Ca	20	●●	●●●●●●●●●●●●●●●●●●●	●●●●●●●●●●●●●●●●●●●●	●●	2,8,8,2

Table 2.10 The electron arrangements of the first 20 elements.

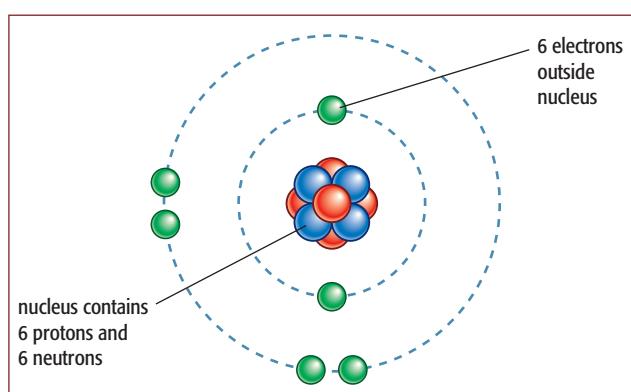


Figure 2.34 Possibly the most versatile atom in the Universe – the carbon-12 atom.



Questions

- 2.20** What are the maximum numbers of electrons that can fill the first and the second shells (energy levels) of an atom?
- 2.21** What is the electron arrangement of a calcium atom, which has an atomic number of 20?
- 2.22** How many electrons are there in the outer shells of the atoms of the noble gases, argon and neon?
- 2.23** Carbon-12 and carbon-14 are different isotopes of carbon. How many electrons are there in an atom of each isotope?

Study tip

Make sure that you remember how to work out the electron arrangements of the first 20 elements and can draw them in rings (shells) as in Figure 2.35. Also remember that you can give the electron arrangement or **electronic structure** simply in terms of numbers: 2,8,4 for silicon, for example.

You can see from these elements that the number of outer electrons in an atom is the same as the number of the group in the Periodic Table that the element is in. The number of shells of electrons in an atom tells you the period (row) of the element in the table. We will look at this further in the next chapter.

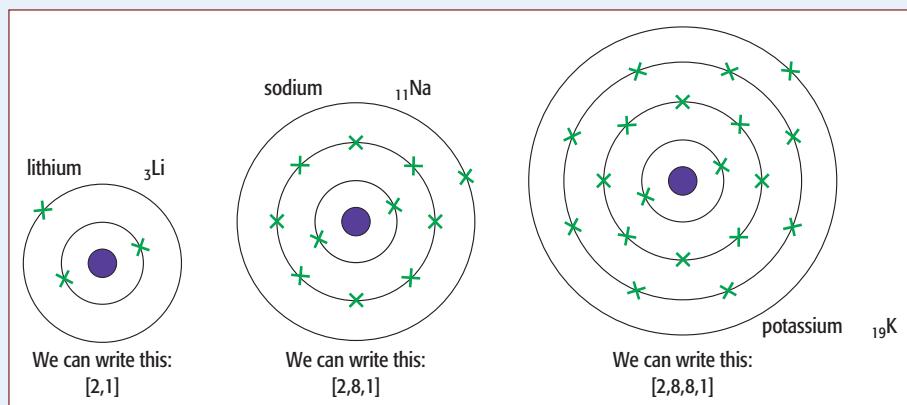


Figure 2.35 Different ways of showing electron structure.

Summary

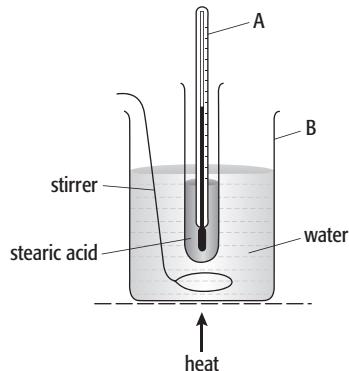
You should know:

- ◆ that there are three different physical states in which a substance can exist
- ◆ about the different changes in state that can take place, including sublimation, where the liquid phase is bypassed
- ◆ how these changes of state can be produced by changing conditions of temperature and/or pressure
- ◆ how the kinetic theory describes the idea that the particles of a substance are in constant motion and that the nature and amount of motion of these particles differs in a solid, liquid or gas
- ◆ how changing physical state involves energy being absorbed or given out, the temperature of the substance staying constant while the change takes place
- ◆ how pure substances have precise melting and boiling points – their sharpness can be taken as an indication of the degree of purity of the substance
- ◆ that different separation methods – such as filtration, distillation and chromatography – can be used to purify a substance from a mixture
- ◆ how pure chemical substances can be either elements or compounds
- ◆ that elements are the basic building units of the material world – they cannot be chemically broken down into anything simpler
- ◆ how compounds are made from two or more elements chemically combined together, and that their properties are very different from those of the elements they are made from

- ◆ how each element is made from atoms and that atoms can join together to make the molecules either of an element or of a compound
- ◆ how the atoms of the elements are made up of different combinations of the subatomic particles – protons, neutrons and electrons
- ◆ the electrical charges and relative masses of these subatomic particles
- ◆ how, in any atom, the protons and neutrons are bound together in a central nucleus, and the electrons ‘orbit’ the nucleus in different energy levels (or shells)
- ◆ that the number of protons in an atom is defined as the proton (atomic) number (Z) of the element
- ◆ that the nucleon (mass) number (A) is defined as the total number of protons and neutrons in any atom
- ◆ how isotopes of the same element can exist which differ only in the number of neutrons in their nuclei
- ◆ how some isotopes of many elements have unstable nuclei and this makes them radioactive
- ◆ that the different forms of radiation from radioisotopes have scientific, industrial and medical uses
- ◆ how the electrons in atoms are arranged in different energy levels that are at different distances from the nucleus of the atom
- ◆ how each energy level has a maximum number of electrons that it can contain and that the electrons fill the shells closest to the nucleus first.

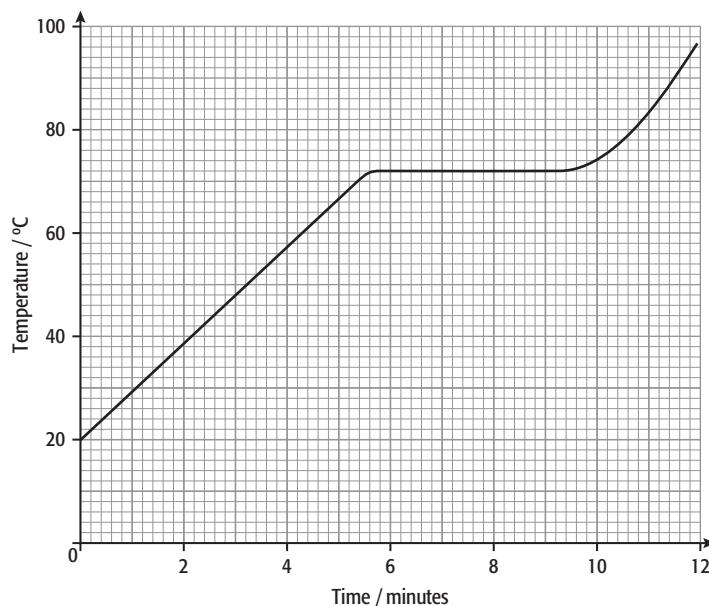
End-of-chapter questions

- 1 Substances can be categorised in two ways: as an element, mixture or compound or as a solid, liquid or gas. Which of these methods is of most use to a chemist?
- 2 The word **particle** can be used to describe a speck of dust, a molecule, an atom or an electron. How can we avoid confusion in using the word particle?
- 3 Stearic acid is a solid at room temperature.
The diagram below shows the apparatus used for finding the melting point of stearic acid. The apparatus was heated at a steady rate and the temperature recorded every minute.



- a State the names of the pieces of apparatus labelled A and B. [2]
- b Suggest why the water needs to be kept stirred during this experiment. [1]

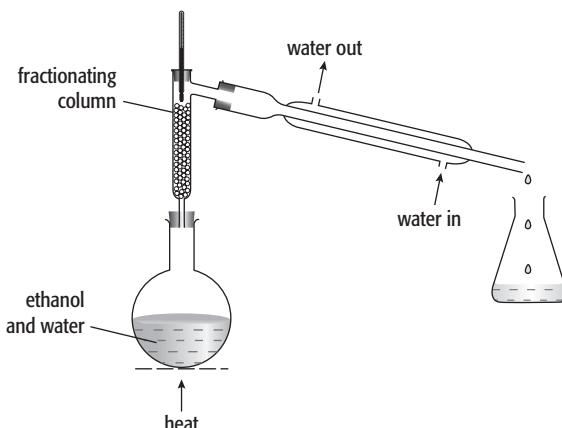
- c A graph of temperature of stearic acid against time of heating is shown below.



- i What was the temperature of the stearic acid after 3 minutes' heating? [1]
- ii Use the information on the graph to determine the melting point of stearic acid. [1]
- d Describe the arrangement and motion of the particles in liquid stearic acid. [2]
- e A sample of stearic acid contained 1% of another compound with a higher relative molecular mass.
 - i Which one of the following statements about this sample of stearic acid is correct?
 - Its density is exactly the same as that of pure stearic acid.
 - Its boiling point is the same as that of pure stearic acid.
 - Its melting point is different from pure stearic acid.
 - Its melting point is the same as that of pure stearic acid. [1]
 - ii Describe one area of everyday life where the purity of substances is important. [1]

[Cambridge IGCSE® Chemistry 0620/21, Question 1(a, b(i), c-e), June 2012]

- 4 Sand and salt (sodium chloride) are both solids.
- a i Describe the arrangement and movement of the particles in a solid. [2]
 - ii Describe how you could separate the sand from a mixture of sand and salt. Give full details of how this is carried out. [3]
 - b The diagram below shows the apparatus used to separate ethanol and water from a mixture of ethanol and water.



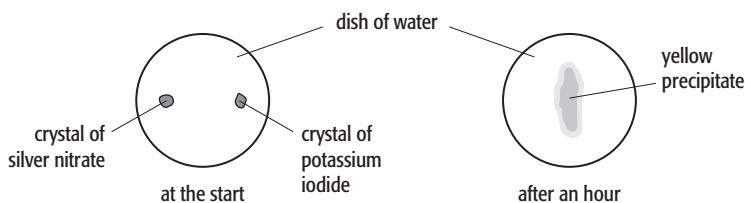
Write out and complete the following sentences about this separation using words from the list below.

condenser	crystallisation	distillation	flask	heavy
higher	lower	solid	volatile	vapour

Fractional is used to separate a mixture of water and ethanol. The temperature at the top of the fractionating column is than the temperature at the bottom. The more liquid evaporates and moves further up the column. It eventually reaches the where the changes to a liquid. [5]

[Cambridge IGCSE® Chemistry 0620/21, Question 3(c, d), November 2012]

- 5 A student placed a crystal of silver nitrate and a crystal of potassium iodide in a dish of water. After an hour she observed that:
- ◆ the crystals had disappeared
 - ◆ a yellow precipitate had appeared near the middle of the dish.



Use your knowledge of the kinetic particle theory and reactions between ions to explain these observations. [4]

[Cambridge IGCSE® Chemistry 0620/21, Question 6(a), November 2012]

- 6 Vanadium has two isotopes.



- a Define the term **isotope**. [1]
b An atom contains protons, electrons and neutrons.

Complete the table to show the number of protons, electrons and neutrons in these two isotopes of vanadium.

Isotope	Number of protons	Number of electrons	Number of neutrons
$^{50}_{23}\text{V}$	23	23	
$^{51}_{23}\text{V}$			28

[3]

- c Write out and complete these sentences using words from the list.

cancer	extra	industry	influenza	medicine	non
--------	-------	----------	-----------	----------	-----

Two types of isotopes are radioactive and -radioactive. Radioactive isotopes are used in for treating patients with [3]

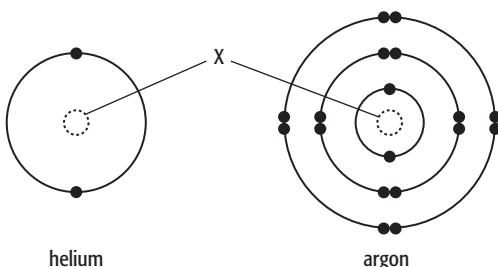
[Cambridge IGCSE® Chemistry 0620/21, Question 2(a-c), June 2011]

7 Helium and argon are noble gases.

a State **one** use of helium.

[1]

b The atomic structures of helium and argon are shown below.



i State the name of the central part of the atom, labelled X.

[1]

ii Which **one** of these statements about helium and argon is correct?

Argon has an incomplete inner shell of electrons.

An atom of argon has 16 electrons.

Helium has a complete outer shell of electrons.

Helium has an incomplete outer shell of electrons.

[1]

iii How many protons are there in an atom of argon?

[1]

iv The symbol for a particular isotope of helium is written as ${}^4_2\text{He}$.

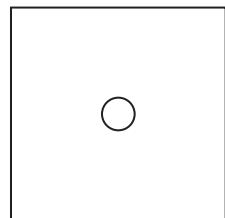
Write a similar symbol for the isotope of argon which has 16 neutrons.

[1]

c Argon is a liquid at a temperature of -188°C .

Complete the diagram below to show how the atoms of argon are arranged at -188°C .

○ represents one atom of argon



[2]

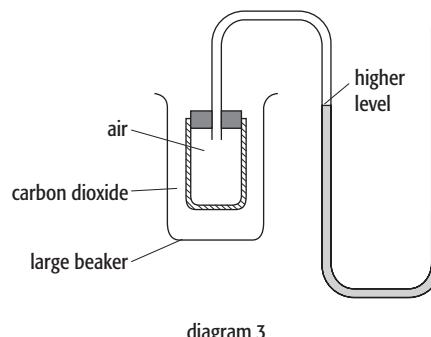
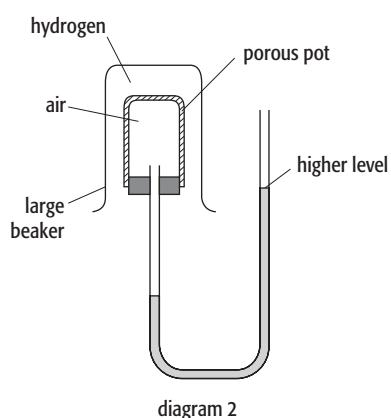
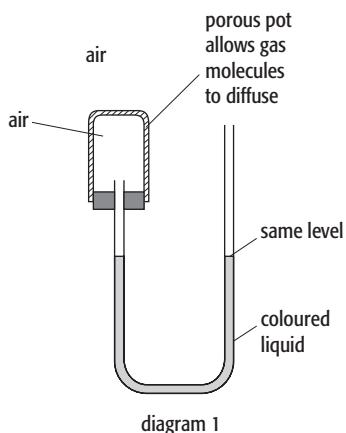
[Cambridge IGCSE® Chemistry 0620/21, Question 3, November 2010]

- S** 8 a A small amount of liquid bromine is added to a container which is then sealed.



Use the ideas of the kinetic theory to explain why, after about an hour, the bromine molecules have spread uniformly to occupy the whole container. [3]

- b The diagrams below show simple experiments on the speed of diffusion of gases.



Write explanations for what is occurring in each diagram. Diagram 1 has been done for you.

Diagram 1

There is air inside and outside the porous pot so the rate of diffusion of air into the pot is the same as the rate of diffusion of air out of the pot. The pressure inside and outside the pot is the same so the coloured liquid is at the same level on each side of the tube.

Diagram 2

[3]

Diagram 3

[3]

[Cambridge IGCSE® Chemistry 0620/33, Question 3, November 2012]

3 Elements and compounds

In this chapter, you will find out about:

- ◆ the structure of the Periodic Table
- ◆ metals and non-metals in the Periodic Table
- ◆ electron arrangement in the Periodic Table
- ◆ trends in Group I – the alkali metals
- ◆ trends in Group VII – the halogens
- ◆ the noble gases
- ◆ trends across a period
- ◆ the transition elements
- ◆ bonding in metals
- ◆ bonding in covalent compounds
- ◆ bonding in ionic compounds
- ◆ formulae and names of ionic compounds
- ◆ formulae and names of covalent compounds
- ◆ the nature of metal crystals and alloys
- ◆ the nature of ionic crystals
- ◆ the nature of giant covalent structures.

Organising the building blocks!

Building up the modern **Periodic Table** has been a major scientific achievement! The first steps towards working out this table were taken long before anyone had any ideas about the structure of atoms. The number of elements discovered increased steadily during the nineteenth century. Chemists began to find patterns in their properties. Döbereiner, Newlands and Meyer all described groupings of elements with similar chemical and physical characteristics. But, although they were partly successful, these groupings were limited or flawed. The breakthrough came in 1869 when Mendeleev put forward his ideas of a periodic table. In his first attempt he used 32 of the 61 elements known at that time (Figure 3.1).

He drew up his table based on atomic masses, as others had done before him. But his success was mainly due to his leaving gaps for possible elements still to be discovered. He did not try to force the elements into patterns for which there was no evidence.

Mendeleev's great achievement lay in predicting the properties of elements that had not yet been discovered.

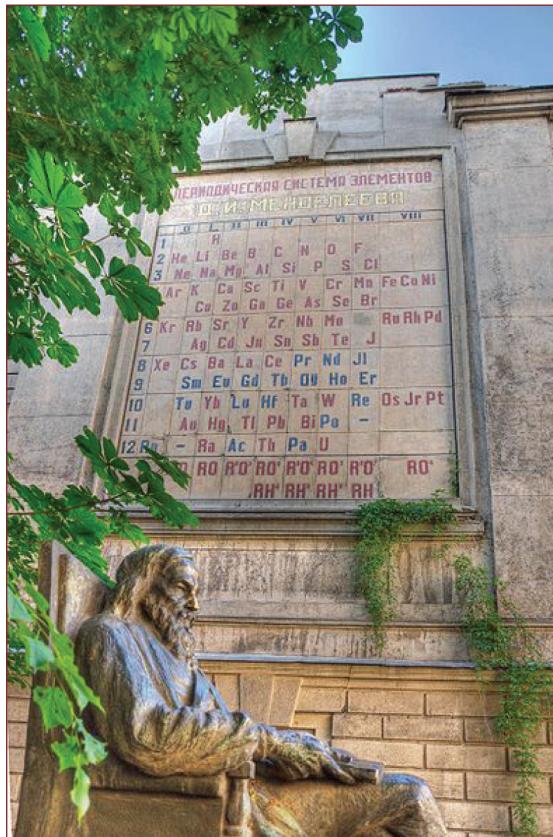


Figure 3.1 Mendeleev's early Periodic Table carved on the wall of a university building in St Petersburg, with a statue of Mendeleev in front.

3.1 The Periodic Table – classifying the elements

All modern versions of the Periodic Table are based on the one put forward by Mendeleev. An example is given in Figure 3.2.

In the Periodic Table:

- ◆ the elements are arranged in order of increasing proton number (atomic number)
- ◆ the vertical columns of elements with similar properties are called **groups**
- ◆ the horizontal rows are called **periods**.

The main distinction in the table is between **metals** and **non-metals**. Metals are clearly separated from non-metals. The non-metals are grouped into the top right-hand region of the table, above the thick

stepped line in Figure 3.2. One of the first uses of the Periodic Table now becomes clear. Although we may never have seen a sample of the element hafnium (Hf), we know from a glance at the table that it is a metal. We may also be able to predict some of its properties.

Metals and non-metals

There are 94 naturally occurring elements. Some are very rare. Francium, for instance, has never been seen. The radioactive metals neptunium and plutonium, which we make artificially in quite large amounts, occur only in very small (trace) quantities naturally. Most of the elements (70) can be classified as metals. Together they form a group of elements whose structures are held together by a particular type of bonding between the atoms. The metals have a number of physical properties that are broadly the same for all of them (Table 3.1).

		Key:																																										
		$\begin{array}{ c } \hline a \\ \hline X \\ \hline b \\ \hline \end{array}$		$a = \text{relative atomic mass}$ $X = \text{symbol}$ $b = \text{atomic number}$																																								
		$\begin{array}{ c } \hline 1 & H \\ \hline \text{Hydrogen} & 1 \\ \hline \end{array}$																																										
	Group I																																											
Period 1																																												
Period 2	7 Li Lithium 3	9 Be Beryllium 4																																										
Period 3	11 Na Sodium 11	12 Mg Magnesium 12																																										
Period 4	19 K Potassium 19	20 Ca Calcium 20	45 Sc Scandium 21	48 Ti Titanium 22	51 V Vanadium 23	52 Cr Chromium 24	55 Mn Manganese 25	56 Fe Iron 26	59 Co Cobalt 27	59 Ni Nickel 28	64 Cu Copper 29	65 Zn Zinc 30	27 Al Aluminium 13	28 Si Silicon 14	11 B Boron 5	12 C Carbon 6	14 N Nitrogen 7	16 O Oxygen 8	19 F Fluorine 9	4 He Helium 2	20 Ne Neon 10																							
Period 5	37 Rb Rubidium 37	38 Sr Strontium 38	39 Y Yttrium 39	40 Zr Zirconium 40	41 Nb Niobium 41	42 Mo Molybdenum 42	43 Tc Technetium 43	44 Ru Ruthenium 44	45 Rh Rhodium 45	46 Pd Palladium 46	47 Ag Silver 47	48 Cd Cadmium 48	49 In Indium 49	50 Sn Tin 51	51 As Arsenic 33	52 Se Selenium 34	53 Br Bromine 35	54 Kr Krypton 36	55 Cl Chlorine 17	56 Ar Argon 18																								
Period 6	55 Cs Caesium 55	56 Ba Barium 56	57 La Cesium 56	58 La Lanthanum 72	59 Hf Hafnium 72	60 Ta Tantalum 73	61 W Tungsten 74	62 Re Rhenium 75	63 Os Osmium 76	64 Ir Iridium 77	65 Pt Platinum 78	66 Au Gold 79	67 Hg Mercury 80	68 Tb Thallium 81	69 Ge Germanium 32	70 As Antimony 50	71 Sb Tellurium 52	72 Te Iodine 53	73 I Xenon 54	74 At Astatine 85	75 Rn Radon 86																							
Period 7	87 Fr Francium 87	88 Ra Radium 88	89 Ac Actinium 88	90 Th Thorium 90	91 Pa Protactinium 91	92 U Uranium 92	93 Np Neptunium 93	94 Pu Plutonium 94	95 Am Americium 95	96 Cm Curium 96	97 Bk Berkelium 97	98 Cf Californium 98	99 Es Einsteinium 99	100 Fm Fermium 100	101 Md Mendelevium 101	102 No Nobelium 102	103 Lr Lawrencium 103	104 Lu Lutetium 71	105 Yb Ytterbium 70	106 Tm Thulium 69	107 Er Erbium 68	108 Ho Holmium 67	109 Dy Dysprosium 66	110 Tb Terbium 65	111 Gd Gadolinium 64	112 Eu Europium 63	113 Sm Samarium 62	114 Nd Neodymium 60	115 Pr Praseodymium 59	116 Ce Cerium 58	117 La Lanthanum 57	118 Ce Curium 58	119 Pr Neptunium 93	120 Nd Plutonium 94	121 Pm Americium 95	122 Sm Curium 96	123 Tb Berkelium 97	124 Gd Californium 98	125 Dy Einsteinium 99	126 Ho Fermium 100	127 Er Mendelevium 101	128 Tm Nobelium 102	129 Yb Lawrencium 103	130 Lu Lutetium 71
Elements in Groups I to 0 are sometimes known as the main-group elements .																																												
 The reactive metals : Group I – the alkali metals; Group II – the alkaline earth metals											 The 'poor' metals																																	
 The transition elements : hard, strong and dense metals											 The non-metals : includes Group VII – the halogens																																	
 The metalloids : includes semiconductors, e.g. silicon and germanium											 The noble gases : very unreactive																																	

Figure 3.2 The Periodic Table, showing the major regions. (Except for chlorine, the relative atomic masses are given to the nearest whole number.)

Metals	Non-metals
They are usually solids (except for mercury, which is a liquid) at room temperature. Their melting and boiling points are usually high.	They are solids or gases (except for bromine, which is a liquid) at room temperature. Their melting and boiling points are often low.
They are usually hard and dense.	Most non-metals are softer than metals (but diamond is very hard). Their densities are often low.
All metals are good conductors of electricity. ^(a)	They are poor conductors of electricity (except graphite, a form of carbon). They tend to be insulators.
They are good conductors of heat.	They are generally poor thermal conductors.
Their shape can be changed by hammering (they are malleable). They can also be pulled out into wires (they are ductile).	Most non-metals are brittle when solid.
They are grey in colour (except gold and copper). They can be polished.	They vary in colour. They often have a dull surface when solid.
They usually make a ringing sound when struck (they are sonorous).	They are not sonorous.

^(a)Electrical conductivity is usually taken as the simplest test of whether a substance is metallic or not.

Table 3.1 Comparison of the physical properties of metals and non-metals.

The chemical properties of metals and non-metals are also very different, as is the type of bonding present in their compounds. The distinction is therefore a very important one.

The Periodic Table does not list substances such as steel, bronze and brass, which in everyday terms we call metals and which share the properties listed for metals. They are **not** elements! They are in fact **alloys**, mixtures of elements (usually metals) designed to have properties that are useful for a particular purpose.

Non-metals are a less uniform group of elements. They show a much wider range of properties. This reflects the wider differences in the types of structure shown by non-metals.

Key definition

- ◆ **metal** – an element that conducts electricity and is malleable and ductile.
- ◆ **non-metal** – is an element that does not conduct electricity well and is neither malleable nor ductile.

Activity 3.1 Testing metals and non-metals

The key test to distinguish between metals and non-metals is electrical conductivity. A simple circuit is set up using either a light bulb or an ammeter. Power is supplied by batteries or a power pack. Examine a range of solid elements and alloys including magnesium, zinc, tin, iron, nickel, roll sulfur, graphite, brass and solder.

A worksheet is included on the CD-ROM.

The change from metallic to non-metallic properties in the elements is not as clear-cut as suggested by drawing the line between the two regions of the Periodic Table. The elements close to the line show properties that lie between these extremes. These elements are now often referred to as **metalloids** (or **semi-metals**). Such elements have some of the properties of metals and others that are more characteristic of non-metals. There are eight elements that are called metalloids. They often look like metals, but are brittle like non-metals. They are

Study tip

If asked to say how you would test to see whether an element was a metal or a non-metal, the key test is **electrical conductivity**. Describe the setting up of a simple circuit using a battery and a light bulb, and then connect in a sample of the element and see if the bulb lights up (Figure 3.3).

The other properties which are most clearly those of a metal are **malleability** and **ductility**. These, and electrical conductivity, are the properties where there are fewest exceptions.

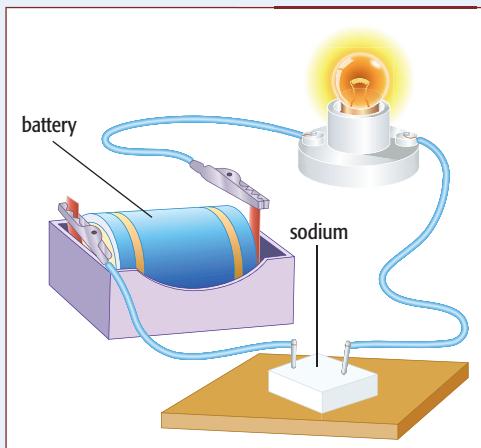


Figure 3.3 Testing the electrical conductivity of a possible metal.



Figure 3.4 A sample of the element silicon, the basis of the semiconductor industry.

neither conductors nor insulators, but make excellent semiconductors. The prime example of this type of element is silicon (Figure 3.4)

Groups and periods in the Periodic Table

The Periodic Table allows us to make even more useful subdivisions of elements than simply deciding which are metals and which are non-metals. The

elements present in Groups I to 0 of the table are sometimes known as the **main-group elements**. These vertical groups show most clearly how elements within the same group have similar chemical and physical properties. Some of these groups have particular names as well as numbers. These are given in Figure 3.2. Between Groups II and III of these main groups of elements is a block of metals known as the **transition elements** (or **transition metals**). The first row of these elements occurs in Period 4. This row includes such important metals as iron, copper and zinc.

The noble gases, in Group 0 on the right-hand side of the table, are the least reactive elements in the table. However, the group next to them, Group VII which are also known as the halogens, and the group on the left-hand side of the table, Group I or the alkali metals, are the most reactive elements. The more unreactive elements, whether metals or non-metals, are in the centre of the table.

Study tip

If you are asked a question about an element in the Periodic Table, use the table at the back of the examination paper to help you answer it.

Electron arrangement and the Periodic Table

When the first attempts were made to construct a Periodic Table, nobody knew about the structure of the atom. We can now directly link the properties of an element with its position in the table and its electron arrangement (Figure 3.5). The number of outer electrons in the atoms of each element has been found. Elements in the same group have the same number of outer electrons. We also know that, as you move across a period in the table, a shell of electrons is being filled.

There is a clear relationship between electron arrangement and position in the Periodic Table for the main-group elements. The elements in Group II have two outer electrons. The elements in Period 3

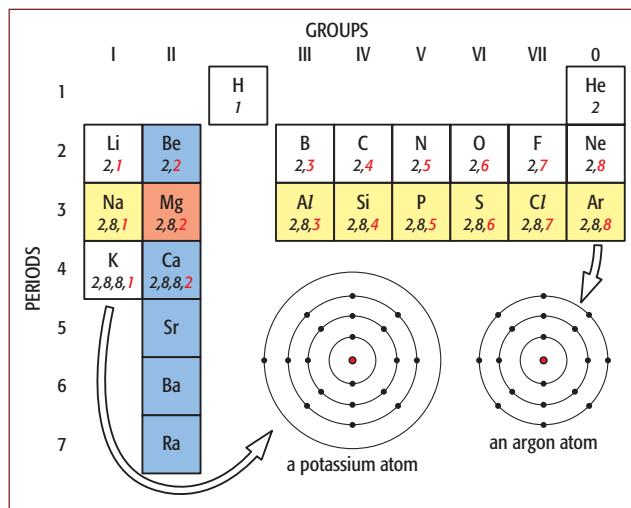


Figure 3.5 The relationship between an element's position in the Periodic Table and the electron arrangement of its atoms.

have three shells of electrons. A magnesium atom has two electrons in its third, outer, shell, and is in Group II. An argon atom has an outer shell containing eight electrons – a very stable arrangement – and is in Group 0. A potassium atom has one electron in its fourth, outer shell, and is in Group I and Period 4.

It is the outer electrons of an atom that are mainly responsible for the chemical properties of any element. Therefore, elements in the same group will have similar properties.

The **electron arrangements** of atoms are linked to position in the Periodic Table.

- ◆ Elements in the same group have the same number of electrons in their outer shell.
- ◆ For the main-group elements, the number of the group is the number of electrons in the outer shell.
- ◆ The periods also have numbers. This number shows us how many shells of electrons the atom has.

Certain electron arrangements are found to be more stable than others. This makes them more difficult to break up. The most stable arrangements are those of the **noble gases**, and this fits in with the fact that they are so unreactive.

There are links between the organisation of particles in the atom and the regular variation in properties of the elements in the Periodic Table. This means that we can see certain broad trends in the table (Figure 3.6). These trends become most obvious if we leave aside the

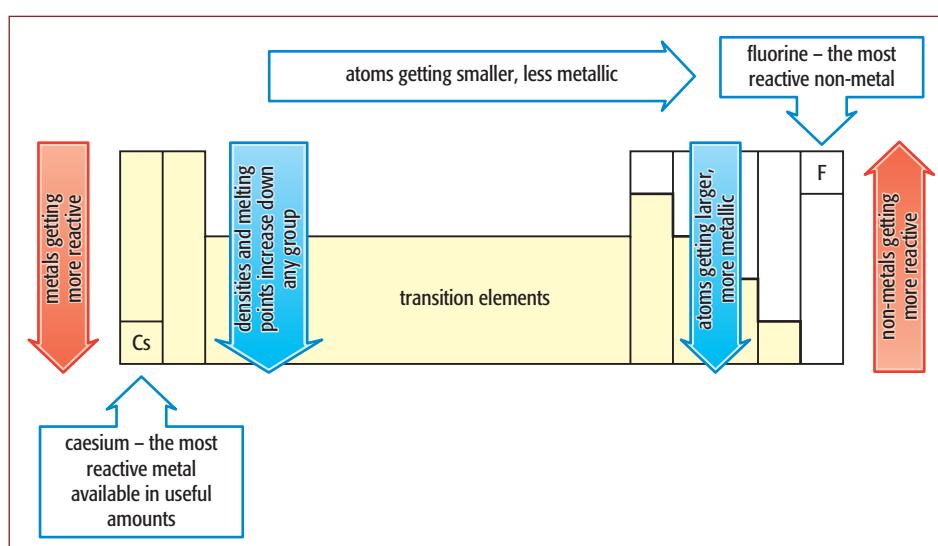


Figure 3.6 General trends in the Periodic Table, leaving aside the noble gases in Group 0.

noble gases in Group 0. Individual groups show certain ‘group characteristics’. These properties follow a trend in particular groups.

Questions

- 3.1 What is the name of the most reactive non-metal?
- 3.2 What is the similarity in the electron arrangement in the noble gases?
- 3.3 How many elements are there in Period 1?
- 3.4 Where, in the Periodic Table, will the largest atom be found?
- 3.5 Sort the following properties into those characteristic of a metal, and those typical of a non-metal.
is an insulator can be beaten into sheets
gives a ringing sound when hit conducts heat
has a dull surface conducts electricity

3.2 Trends in groups

Group I – the alkali metals

The metals in Group I are often called the **alkali metals**. They are soft solids with relatively low melting points and low densities (Figure 3.7). They are highly reactive and are stored in oil to prevent them reacting with the oxygen and water vapour in the air. When freshly cut with a knife, all these metals have a light-grey, silvery surface, which quickly tarnishes (becomes dull). Reactivity increases as we go down the group. All Group I metals react with water to form hydrogen and an alkaline solution of the metal hydroxide. The reactions range from vigorous in the case of lithium



Figure 3.7 The alkali metals are all soft and can be cut with a knife. This is a sample of lithium.

to explosive in the case of caesium. You might predict that francium, at the bottom of Group I, would be the most reactive of all the metals. However, it is highly radioactive and very rare because it decays with a half-life of 5 minutes. It has been estimated that there are only 17 atoms of francium in existence on Earth at any one moment in time.

The physical properties of the alkali metals also change as we go down the group. The melting points become lower while the density of the metals increases.

The **alkali metals** (Group I) are the most reactive metals that occur. They are known as the alkali metals because they react vigorously with water to produce hydrogen and an alkaline solution.

Group VII – the halogens

The most reactive non-metals are the **halogens** in Group VII of the table (Figure 3.8). In contrast with Group I, here reactivity decreases **down** the group. For example, fluorine is a dangerously reactive, pale yellow gas at room temperature. There is a steady increase in melting points and boiling points as we go

Chlorine (Cl_2)
<ul style="list-style-type: none">• dense pale-green gas• smelly and poisonous• occurs as chlorides, especially sodium chloride in the sea• relative atomic mass 35.5
Bromine (Br_2)
<ul style="list-style-type: none">• deep-red liquid with red-brown vapour• smelly and poisonous• occurs as bromides, especially magnesium bromide in the sea• relative atomic mass 80
Iodine (I_2)
<ul style="list-style-type: none">• grey solid with purple vapour• smelly and poisonous• occurs as iodides and iodates in some rocks and in seaweed• relative atomic mass 127

Figure 3.8 The general properties of some of the halogens (Group VII).

down the group, and the elements change from gases to solids as the atomic number increases. Interestingly, the lowest element in this group is also a highly radioactive and rare element, astatine. The actual properties of astatine remain a mystery to us, but we could make a good guess at some of them.

The halogen family found in Group VII of the Periodic Table shows clearly the similarities of elements in the group.

Common properties of the halogens

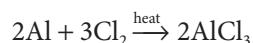
- ◆ They are all poisonous and have a similar strong smell.
 - ◆ They are all non-metals.
 - ◆ They all form diatomic molecules (for example Cl_2 , Br_2 , I_2).
 - ◆ They all have a **valency** (combining power) of 1 and form compounds with similar formulae, for example hydrogen chloride (HCl), hydrogen bromide (HBr), hydrogen iodide (HI).
 - ◆ Their compounds with hydrogen are usually strong acids when dissolved in water, for example hydrochloric acid (HCl), hydrobromic acid (HBr), hydriodic acid (HI).
 - ◆ They each produce a series of compounds with other elements: chlorides, bromides and iodides. Together these are known as **halides**.
 - ◆ The halogens themselves can react directly with metals to form metal halides (or **salts**).
 - ◆ They all form negative ions carrying a single charge, for example chloride ions (Cl^-), bromide ions (Br^-), iodide ions (I^-).

There are gradual changes in properties between the halogens (see Figure 3.8). As you go down the group, the boiling points increase. Also there is a change from gas to liquid to solid. The intensity of the colour of the element also increases, from pale to dark. Following these trends, it should not surprise you to know that fluorine is a pale yellow gas at room temperature.

The chemical reactivity of the halogens

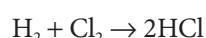
Fluorine and chlorine are very reactive. They combine strongly with both metals and non-metals. A piece of Dutch metal foil – an alloy of copper and zinc – will burst into flames when placed in a gas jar of chlorine.

When chlorine is passed over heated aluminium, the metal glows white and forms aluminium chloride:



Aluminium also reacts strongly with bromine and iodine. The reaction between a dry mixture of powdered aluminium and iodine can be triggered by adding just a few drops of water. The reaction is highly exothermic and some of the iodine is given off as purple fumes before it has a chance to react.

Hydrogen will burn in chlorine to form hydrogen chloride. Carried out a different way, the reaction can be explosive:



Chlorine dissolves in water to give an acidic solution. This mixture is called **chlorine water** and contains two acids:



Chlorine water acts as an **oxidising agent** – hypochlorous acid can give up its oxygen to other substances. It also acts as a bleach because some coloured substances lose their colour when they are oxidised. This reaction is used as the chemical test for chlorine gas. Damp litmus or **Universal Indicator** paper is bleached when held in the gas. The halogens become steadily less reactive as you go down the group. Table 3.3 (overleaf) gives some examples of the reactivity of the halogens.

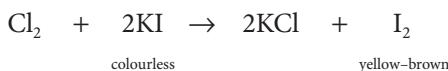
The **displacement reactions** shown in the lower part of Table 3.2 demonstrate the order of reactivity of the three major halogens. For example, if you add chlorine to a solution of potassium bromide, the chlorine displaces bromine (Figure 3.9, overleaf). Chlorine is more reactive than bromine, so it replaces it and potassium chloride is formed. Potassium bromide solution is colourless. It turns orange when chlorine is bubbled through it:



Reaction with	Chlorine	Bromine	Iodine
coloured dyes	bleaches easily	bleaches slowly	bleaches very slowly
iron wool	iron wool reacts strongly to form iron(III) chloride; needs heat to start	iron reacts steadily to form iron(III) bromide; needs continuous heating	iron reacts slowly, even with continuous heating, to form iron(III) iodide
chlorides	—	no reaction	no reaction
bromides	displaces bromine, e.g. $\text{Cl}_2 + 2\text{KBr} \rightarrow 2\text{KCl} + \text{Br}_2$	—	no reaction
iodides	displaces iodine, e.g. $\text{Cl}_2 + 2\text{KI} \rightarrow 2\text{KCl} + \text{I}_2$	displaces iodine, e.g. $\text{Br}_2 + 2\text{KI} \rightarrow 2\text{KBr} + \text{I}_2$	—

Table 3.2 Some reactions of the halogens.

Chlorine will also displace iodine from potassium iodide:



You will find more information about the halogens and their uses in Chapter 9, Table 9.3.

Study tip

If you are asked to put elements from a group in order of reactivity, you must be very careful when reading the question to see whether the answer should be in order of increasing or decreasing reactivity.

Group 0 – the noble gases

When Mendeleev first constructed his table, part of his triumph was to predict the existence and properties of some undiscovered elements. However, there was no indication that a whole **group** of elements (Group 0) remained to be discovered! Because of their lack of reactivity, there was no clear sign of their existence.

However, analysis of the gases in air led to the discovery of argon. There was no suitable place in the table for an individual element with argon's properties. This pointed to the existence of an entirely new group! In the 1890s, helium, which had first been detected by spectroscopy of light from the Sun during an eclipse, and the other noble gases in the group (Group 0) were isolated.

The radioactive gas, radon, was the last to be purified,

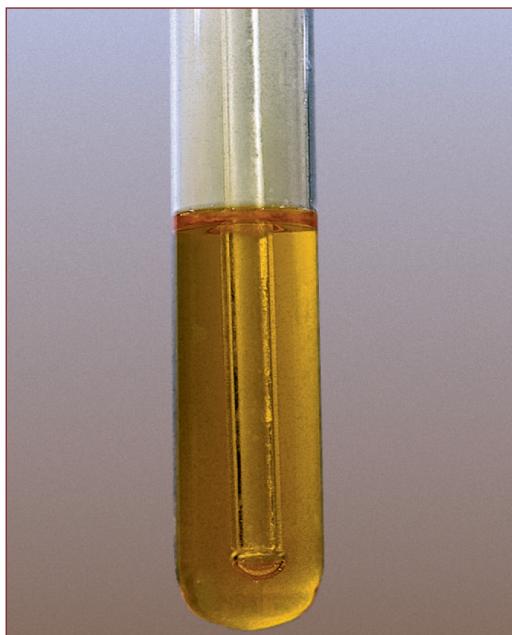


Figure 3.9 Bromine is displaced by chlorine from a colourless solution of potassium bromide.

in 1908. One man, William Ramsay, was involved in the isolation of all the elements in the group. He was awarded the Nobel Prize for this major contribution.

All the noble gases are present in the Earth's atmosphere. Together they make up about 1% of the total, though argon is the most common. These gases are particularly unreactive. They were sometimes referred to as the **inert gases**, meaning they did not react at all. However, since the 1960s, some compounds of xenon and krypton have been made and their name was changed to the **noble gases**. The uses of