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# Chemistry

Third Edition

Bryan Earl  
Doug Wilford

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## International hazard warning symbols

You will need to be familiar with these symbols when undertaking practical experiments in the laboratory.



### Corrosive

These substances attack or destroy living tissues, including eyes and skin.



### Oxidising

These substances provide oxygen which allows other materials to burn more fiercely.



### Harmful

These substances are similar to toxic substances but less dangerous.



### Toxic

These substances can cause death.



### Irritant

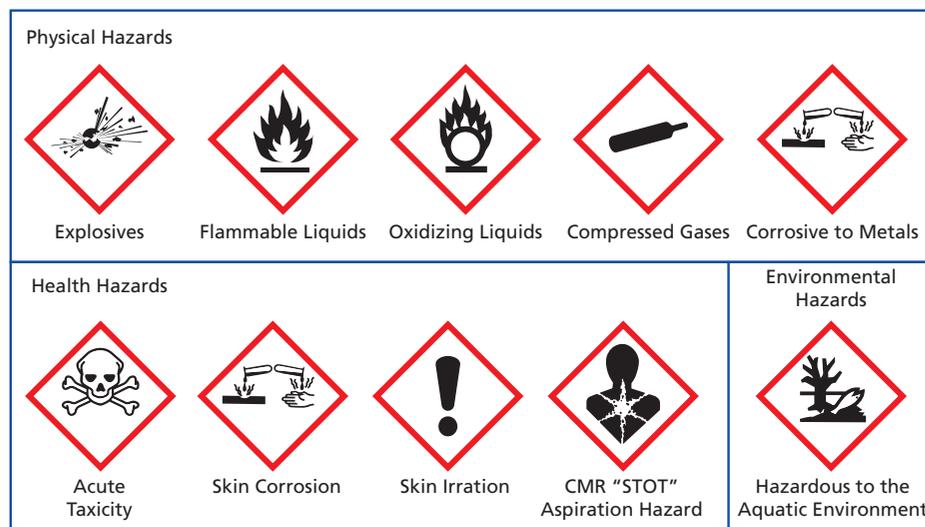
These substances are not corrosive but can cause red- dening or blistering of the skin.



### Highly flammable

These substances can easily catch fire.

Teachers and students should note that a new system for labelling hazards is being introduced between 2010 and 2015 and, in due course, you will need to become familiar with these new symbols:



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# Contents

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	Acknowledgements	vii
	Preface to the reader	ix
<b>Chapter 1</b>	<b>The particulate nature of matter</b>	<b>1</b>
	Solids, liquids and gases	2
	The kinetic theory of matter	2
	Changes of state	4
	Diffusion – evidence for moving particles	6
	Checklist	8
	Additional questions	9
<b>Chapter 2</b>	<b>Elements, compounds and experimental techniques</b>	<b>10</b>
	Elements	10
	Compounds	13
	Mixtures	16
	Separating mixtures	17
	Accuracy in experimental work in the laboratory	25
	Gels, sols, foams and emulsions	26
	Mixtures for strength	28
	Checklist	29
	Additional questions	31
<b>Chapter 3</b>	<b>Atomic structure and bonding</b>	<b>33</b>
	Inside atoms	33
	The arrangement of electrons in atoms	37
	Ionic bonding	38
	Covalent bonding	45
	Glasses and ceramics	54
	Metallic bonding	55
	Checklist	56
	Additional questions	58
<b>Chapter 4</b>	<b>Stoichiometry – chemical calculations</b>	<b>59</b>
	Relative atomic mass	59
	Reacting masses	59
	Calculating moles	61
	Calculating formulae	64
	Moles and chemical equations	66
	Checklist	69
	Additional questions	71

<b>Chapter 5</b>	<b>Electricity and chemistry</b>	<b>72</b>
	Electrolysis of lead(II) bromide	73
	Electrolysis of aluminium oxide	74
	Electrolysis of aqueous solutions	77
	Electrolysis of concentrated hydrochloric acid	80
	Electrolysis of copper(II) sulfate solution	80
	Electrolysis guidelines	83
	Electroplating	83
	Checklist	85
	Additional questions	86
<b>Chapter 6</b>	<b>Chemical energetics</b>	<b>88</b>
	Substances from oil	88
	Fossil fuels	90
	What is a fuel?	92
	Alternative sources of energy	93
	Chemical energy	95
	Changes of state	97
	Cells and batteries	98
	Checklist	100
	Additional questions	101
<b>Chapter 7</b>	<b>Chemical reactions</b>	<b>104</b>
	Factors that affect the rate of a reaction	105
	Enzymes	111
	Checklist	114
	Additional questions	115
<b>Chapter 8</b>	<b>Acids, bases and salts</b>	<b>117</b>
	Acids and alkalis	117
	Formation of salts	122
	Crystal hydrates	127
	Solubility of salts in water	129
	Titration	129
	Checklist	132
	Additional questions	133
<b>Chapter 9</b>	<b>The Periodic Table</b>	<b>135</b>
	Development of the Periodic Table	135
	Electronic structure and the Periodic Table	138
	Group I – the alkali metals	138
	Group II – the alkaline earth metals	140
	Group VII – the halogens	141
	Group 0 – the noble gases	143
	Transition elements	144
	The position of hydrogen	146
	Checklist	146
	Additional questions	147

<b>Chapter 10</b>	<b>Metals</b>	<b>149</b>
	Metal reactions	150
	Decomposition of metal nitrates, carbonates, oxides and hydroxides	152
	Reactivity of metals and their uses	153
	Identifying metal ions	155
	Discovery of metals and their extraction	157
	Metal waste	161
	Rusting of iron	161
	Alloys	165
	Checklist	168
	Additional questions	169
<b>Chapter 11</b>	<b>Air and water</b>	<b>171</b>
	The air	171
	How do we get the useful gases we need from the air?	174
	Ammonia – an important nitrogen-containing chemical	176
	Artificial fertilisers	180
	Atmospheric pollution	182
	Water	184
	The water cycle	186
	Hardness in water	187
	Water pollution and treatment	190
	Checklist	193
	Additional questions	194
<b>Chapter 12</b>	<b>Sulfur</b>	<b>197</b>
	Sulfur – the element	197
	Sulfur dioxide	198
	Sulfuric acid	199
	Checklist	203
	Additional questions	204
<b>Chapter 13</b>	<b>Inorganic carbon chemistry</b>	<b>206</b>
	Limestone	206
	Carbonates	211
	Carbon dioxide	212
	Checklist	215
	Additional questions	216
<b>Chapter 14</b>	<b>Organic chemistry 1</b>	<b>218</b>
	Alkanes	218
	The chemical behaviour of alkanes	220
	Alkenes	222
	The chemical behaviour of alkenes	224
	A special addition reaction of alkene molecules	226
	Checklist	230
	Additional questions	231

<b>Chapter 15</b>	<b>Organic chemistry 2</b>	<b>233</b>
	Alcohols (R—OH)	233
	Biotechnology	236
	Carboxylic acids	237
	Soaps and detergents	239
	Condensation polymers	241
	Some biopolymers	242
	Pharmaceuticals	246
	Checklist	247
	Additional questions	249
<b>Chapter 16</b>	<b>Experimental chemistry</b>	<b>251</b>
	Objectives for experimental skills and investigations	251
	Suggestions for practical work and assessment	251
	Notes on qualitative analysis	261
	<b>Revision and exam-style questions</b>	<b>264</b>
	Alternative to practical paper	264
	Theory	274
	<b>The Periodic Table of the elements</b>	<b>294</b>
	<b>Index</b>	<b>295</b>

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# Preface to the reader

---

This textbook has been written to help you in your study of chemistry to Cambridge IGCSE. The different chapters in this book are split up into short topics. At the end of many of these topics are questions to test whether you have understood what you have read. At the end of each chapter there are larger study questions. Try to answer as many of the questions as you can as you come across them because asking and answering questions is at the heart of your study of chemistry.

Some questions in the style of Cambridge IGCSE examination papers are included at the end of the book. In many cases they are designed to test your ability to apply your chemical knowledge. The questions may provide certain facts and ask you to make an interpretation of them. In such cases, the factual information may not be covered in the text.

To help draw attention to the more important words, scientific terms are printed in **bold** the first time they are used. There are also checklists at the end of each chapter summarising the important points covered.

As you read through the book, you will notice three sorts of shaded area in the text.

Material highlighted in green is for the Cambridge IGCSE Extended curriculum.

Areas highlighted in yellow contain material that is not part of the Cambridge IGCSE syllabus. It is extension work and will not be examined.

Questions are highlighted by a box like this.

We use different colours to define different areas of chemistry:

- 'starter' chapters – basic principles
- physical chemistry
- inorganic chemistry
- organic chemistry and the living world.

You will see from the box at the foot of this page that the book is divided into four different areas of chemistry: Starter, Physical, Inorganic and Organic chemistry. We feel, however, that some topics lead naturally on to other topics not in the same area. So you can, of course, read and study the chapters in your own preferred order and the colour coding will help you with this.

The accompanying **Revision CD-ROM** provides invaluable exam preparation and practice. We want to test your knowledge with interactive questions that cover both the Core and Extended curriculum. These are organised by syllabus topic.

Together, the textbook and CD-ROM will provide you with the information you need for the Cambridge IGCSE syllabus. We hope you enjoy using them.

Bryan Earl and Doug Wilford

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## 1

# The particulate nature of matter

## Solids, liquids and gases

### The kinetic theory of matter

Explaining the states of matter

### Changes of state

An unusual state of matter

An unusual change of state

Heating and cooling curves

## Diffusion – evidence for moving particles

Brownian motion

## Checklist

## Additional questions

Chemistry is about what **matter** is like and how it behaves, and our explanations and predictions of its behaviour. What is matter? This word is used to cover all the substances and materials from which the physical universe is composed. There are many millions of different substances known, and all of them can be categorised as solids, liquids or gases (Figure 1.1). These are what we call the **three states of matter**.



a solid



b liquid



c gas

**Figure 1.1** Water in three different states.

## ● Solids, liquids and gases

A **solid**, at a given temperature, has a definite volume and shape which may be affected by changes in temperature. Solids usually increase slightly in size when heated (**expansion**) (Figure 1.2) and usually decrease in size if cooled (**contraction**).

A **liquid**, at a given temperature, has a fixed volume and will take up the shape of any container into which it is poured. Like a solid, a liquid's volume is slightly affected by changes in temperature.

A **gas**, at a given temperature, has neither a definite shape nor a definite volume. It will take up the shape of any container into which it is placed and will spread out evenly within it. Unlike those of solids and liquids, the volumes of gases are affected quite markedly by changes in temperature.

Liquids and gases, unlike solids, are relatively **compressible**. This means that their volume can be reduced by the application of pressure. Gases are much more compressible than liquids.



**Figure 1.2** Without expansion gaps between the rails, the track would buckle in hot weather.

## ● The kinetic theory of matter

The **kinetic theory** helps to explain the way in which matter behaves. The evidence is consistent with the idea that all matter is made up of tiny **particles**. This theory explains the physical properties of matter in terms of the movement of its constituent particles.

The main points of the theory are:

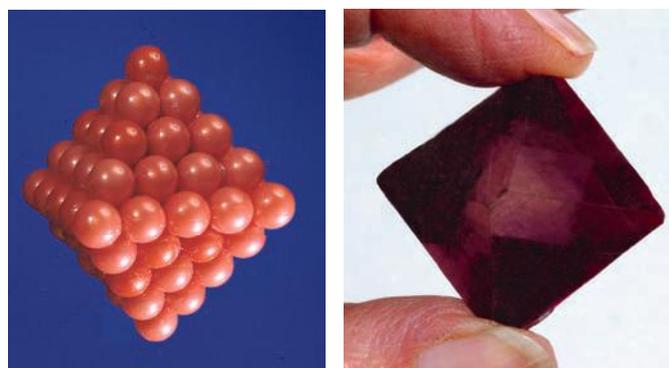
- All matter is made up of tiny, moving particles, invisible to the naked eye. Different substances have different types of particles (atoms, molecules or ions) which have different sizes.
- The particles move all the time. The higher the temperature, the faster they move on average.
- Heavier particles move more slowly than lighter ones at a given temperature.

The kinetic theory can be used as a scientific model to explain how the arrangement of particles relates to the properties of the three states of matter.

### Explaining the states of matter

In a solid the particles attract one another. There are attractive forces between the particles which hold them close together. The particles have little freedom of movement and can only vibrate about a fixed position. They are arranged in a regular manner, which explains why many solids form crystals.

It is possible to model such crystals by using spheres to represent the particles (Figure 1.3a). If the spheres are built up in a regular way then the shape compares very closely with that of a part of a chrome alum crystal (Figure 1.3b).



**a** A model of a chrome alum crystal. **b** An actual chrome alum crystal.

**Figure 1.3**

Studies using X-ray crystallography (Figure 1.4) have confirmed how the particles are arranged in crystal structures. When crystals of a pure substance form under a given set of conditions, the particles present are always packed in the same way. However, the particles may be packed in different ways in crystals of different substances. For example, common salt (sodium chloride) has its particles arranged to give cubic crystals as shown in Figure 1.5.



**Figure 1.4** A modern X-ray crystallography instrument, used for studying crystal structure.

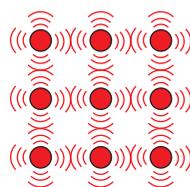


**Figure 1.5** Sodium chloride crystals.

In a liquid the particles are still close together but they move around in a random way and often collide with one another. The forces of attraction between the particles in a liquid are weaker than those in a solid. Particles in the liquid form of a substance have more energy on average than the particles in the solid form of the same substance.

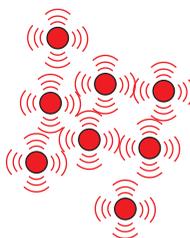
In a gas the particles are relatively far apart. They are free to move anywhere within the container in which they are held. They move randomly at very high velocities, much more rapidly than those in a liquid. They collide with each other, but less often than in a liquid, and they also collide with the walls of the container. They exert virtually no forces of attraction on each other because they are relatively far apart. Such forces, however, are very significant. If they did not exist we could not have solids or liquids (see Changes of state, p. 4).

The arrangement of particles in solids, liquids and gases is shown in Figure 1.6.



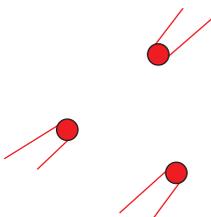
**solid**

Particles only vibrate about fixed positions. Regular structure.



**liquid**

Particles have some freedom and can move around each other. Collide often.



**gas**

Particles move freely and at random in all the space available. Collide less often than in liquid.

**Figure 1.6** The arrangement of particles in solids, liquids and gases.

### Questions

- 1 When a metal such as copper is heated it expands. Explain what happens to the metal particles as the solid metal expands.
- 2 Use your research skills on the Internet to find out about the technique of X-ray crystallography and how this technique can be used to determine the crystalline structure of solid substances such as sodium chloride.

## ● Changes of state

The kinetic theory model can be used to explain how a substance changes from one state to another. If a solid is heated the particles vibrate faster as they gain energy. This makes them ‘push’ their neighbouring particles further away from themselves. This causes an increase in the volume of the solid, and the solid expands. Expansion has taken place.

Eventually, the heat energy causes the forces of attraction to weaken. The regular pattern of the structure breaks down. The particles can now move around each other. The solid has melted. The temperature at which this takes place is called the **melting point** of the substance. The temperature of a pure melting solid will not rise until it has all melted. When the substance has become a liquid there are still very significant forces of attraction between the particles, which is why it is a liquid and not a gas.

Solids which have high melting points have stronger forces of attraction between their particles than those which have low melting points. A list of some substances with their corresponding melting and boiling points is shown in Table 1.1.

Table 1.1

Substance	Melting point/°C	Boiling point/°C
Aluminium	661	2467
Ethanol	-117	79
Magnesium oxide	827	3627
Mercury	-30	357
Methane	-182	-164
Oxygen	-218	-183
Sodium chloride	801	1413
Sulfur	113	445
Water	0	100

If the liquid is heated the particles will move around even faster as their average energy increases. Some particles at the surface of the liquid have enough energy to overcome the forces of attraction between themselves and the other particles in the liquid and they escape to form a gas. The liquid begins to **evaporate** as a gas is formed.

Eventually, a temperature is reached at which the particles are trying to escape from the liquid so quickly that bubbles of gas actually start to form inside the bulk of the liquid. This temperature is

called the **boiling point** of the substance. At the boiling point the pressure of the gas created above the liquid equals that in the air – **atmospheric pressure**.

Liquids with high boiling points have stronger forces between their particles than liquids with low boiling points.

When a gas is cooled the average energy of the particles decreases and the particles move closer together. The forces of attraction between the particles now become significant and cause the gas to **condense** into a liquid. When a liquid is cooled it **freezes** to form a solid. In each of these changes energy is given out.

Changes of state are examples of **physical changes**. Whenever a physical change of state occurs, the temperature remains constant during the change (see Heating and cooling curves, p. 5). During a physical change no new substance is formed.

### An unusual state of matter

**Liquid crystals** are an unusual state of matter (Figure 1.7). These substances look like liquids and flow like liquids but have some order in the arrangement of the particles, and so in some ways they behave like crystals.

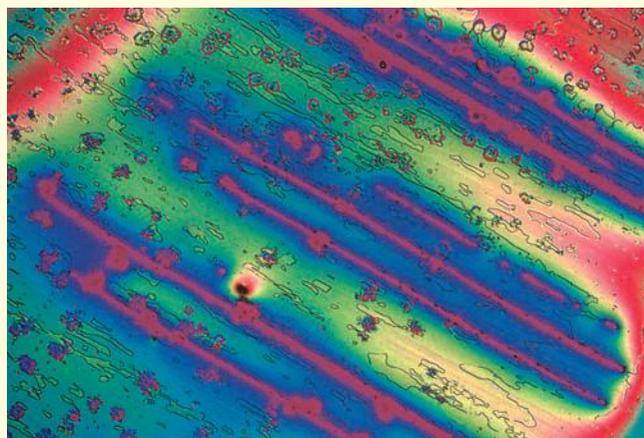


Figure 1.7 A polarised light micrograph of liquid crystals.

Liquid crystals are now part of our everyday life. They are widely used in displays for digital watches, calculators and lap-top computers, and in televisions (Figure 1.8). They are also useful in thermometers because liquid crystals change colour as the temperature rises and falls.



Figure 1.8 Liquid crystals are used in this TV screen.

## An unusual change of state

There are a few substances that change directly from a solid to a gas when they are heated without ever becoming a liquid. This rapid spreading out of the particles is called **sublimation**. Cooling causes a change from a gas directly back to a solid. Examples of substances that behave in this way are carbon dioxide (Figure 1.9) and iodine.

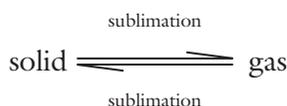


Figure 1.9 Dry ice (solid carbon dioxide) sublimates on heating and can be used to create special effects on stage.

Carbon dioxide is a white solid called dry ice at temperatures below  $-78\text{ }^{\circ}\text{C}$ . When heated to just above  $-78\text{ }^{\circ}\text{C}$  it changes into carbon dioxide gas. The changes of state are summarised in Figure 1.10.

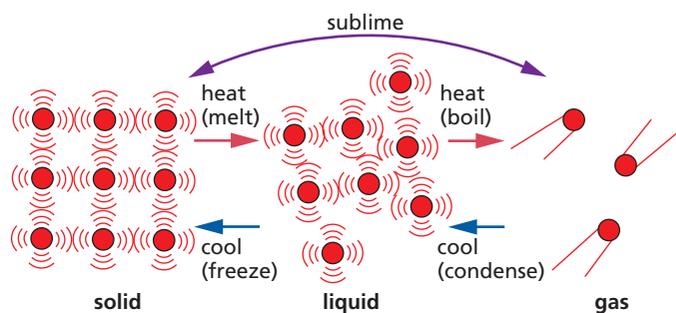


Figure 1.10 Summary of the changes of state.

## Heating and cooling curves

The graph shown in Figure 1.11 was drawn by plotting the temperature of water as it was heated steadily from  $-15\text{ }^{\circ}\text{C}$  to  $110\text{ }^{\circ}\text{C}$ . You can see from the curve that changes of state have taken place. When the temperature was first measured only ice was present. After a short time the curve flattens, showing that even though heat energy is being put in, the temperature remains constant.

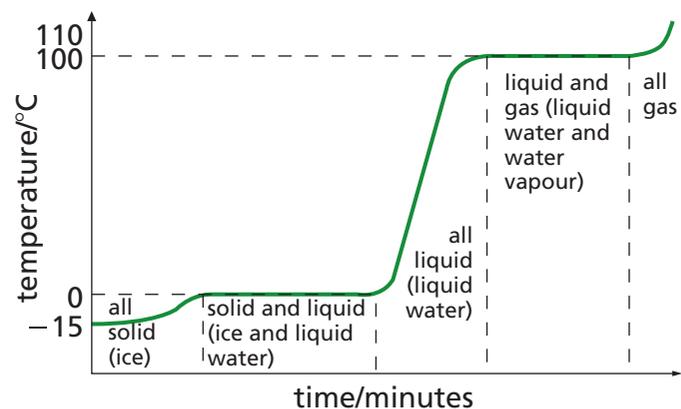
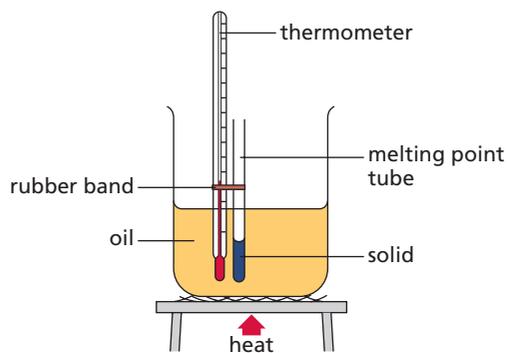


Figure 1.11 Graph of temperature against time for the change from ice at  $-15\text{ }^{\circ}\text{C}$  to water to steam.

In ice the particles of water are close together and are attracted to one another. For ice to melt the particles must obtain sufficient energy to overcome the forces of attraction between the water particles to allow relative movement to take place. This is where the heat energy is going.

The temperature will begin to rise again only after all the ice has melted. Generally, the heating curve for a pure solid always stops rising at its melting point and gives rise to a sharp melting point. A sharp melting point indicates a pure sample. The addition or presence of impurities lowers the melting point. You can try to find the melting point of a substance using the apparatus shown in Figure 1.12.



**Figure 1.12** Apparatus shown here if heated slowly can be used to find the melting point of a substance such as the solid in the melting point tube.

In the same way, if you want to boil a liquid such as water you have to give it some extra energy. This can be seen on the graph (Figure 1.11) where the curve levels out at  $100^{\circ}\text{C}$  – the boiling point of water.

Solids and liquids can be identified from their characteristic melting and boiling points.

The reverse processes of condensing and freezing occur on cooling. This time, however, energy is given out when the gas condenses to the liquid and the liquid freezes to give the solid.

### Questions

- 1 Write down as many uses as you can for liquid crystals.
- 2 Why do gases expand more than solids for the same increase in temperature?
- 3 Ice on a car windscreen will disappear as you drive along, even without the heater on. Explain why this happens.
- 4 When salt is placed on ice the ice melts. Explain why.
- 5 Draw and label the graph you would expect to produce if water at  $100^{\circ}\text{C}$  was allowed to cool to  $-5^{\circ}\text{C}$ .

## ● Diffusion – evidence for moving particles

When you walk past a cosmetics counter in a department store you can usually smell the perfumes. For this to happen gas particles must be leaving open perfume bottles and be spreading out through the air in the store. This spreading out of a gas is called **diffusion** and it takes place in a haphazard and random way.

All gases diffuse to fill the space available. In Figure 1.13, after a day the brown–red fumes of gaseous bromine have spread evenly throughout both gas jars from the liquid present in the lower gas jar.

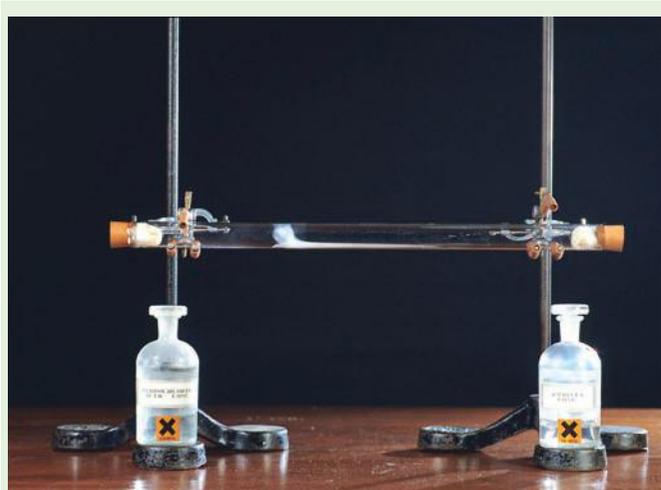


**Figure 1.13** After 24 hours the bromine fumes have diffused throughout both gas jars.

Gases diffuse at different rates. If one piece of cotton wool is soaked in concentrated ammonia solution and another is soaked in concentrated hydrochloric acid and these are put at opposite ends of a dry glass tube, then after a few minutes a white cloud of ammonium chloride appears (Figure 1.14). This shows the position at which the two gases meet and react. The white cloud forms in the position shown because the ammonia particles are lighter and have a smaller relative molecular mass (Chapter 4, p. 62) than the hydrogen chloride particles (released from the hydrochloric acid) and so move faster.

Diffusion also takes place in liquids (Figure 1.15) but it is a much slower process than in gases. This is because the particles of a liquid move much more slowly.

When diffusion takes place between a liquid and a gas it is known as **intimate mixing**. The kinetic theory can be used to explain this process. It states that collisions are taking place randomly between particles in a liquid or a gas and that there is sufficient space between the particles of one substance for the particles of the other substance to move into.



**Figure 1.14** Hydrochloric acid (left) and ammonia (right) diffuse at different rates.



**a**



**b**

**Figure 1.15** Diffusion within nickel(II) sulfate solution can take days to reach the stage shown on the right.

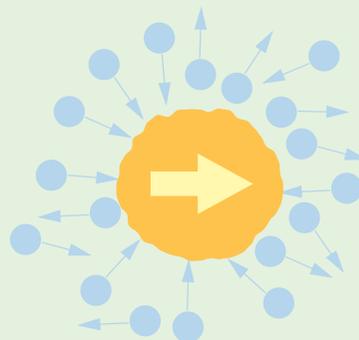
## Questions

- 1 When a jar of coffee is opened, people in all parts of the room soon notice the smell. Use the kinetic theory to explain how this happens.
- 2 Describe, with the aid of diagrams, the diffusion of nickel(II) sulfate solution.
- 3 Explain why diffusion is faster in gases than in liquids.

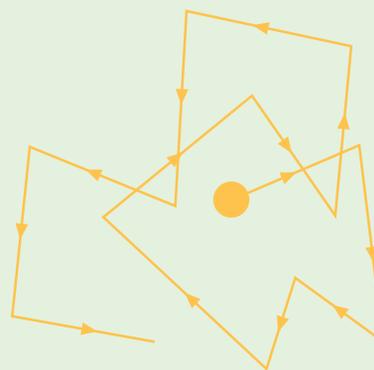
## Brownian motion

Evidence for the movement of particles in liquids came to light in 1827 when a botanist, Robert Brown, observed that fine pollen grains on the surface of water were not stationary. Through his microscope he noticed that the grains were moving about in a random way. It was 96 years later, in 1923, that another scientist called Norbert Wiener explained what Brown had observed. He said that the pollen grains were moving because the much smaller and faster-moving water particles were constantly colliding with them (Figure 1.16a).

This random motion of visible particles (pollen grains) caused by much smaller, invisible ones (water particles) is called **Brownian motion** (Figure 1.16b), after the scientist who first observed this phenomenon. It was used as evidence for the kinetic particle model of matter (p. 3).



**Figure 1.16a** Pollen particle being bombarded by water molecules.



**Figure 1.16b** Brownian motion causes the random motion of the visible particle.

### Checklist

After studying Chapter 1 you should know and understand the following terms.

- **Atmospheric pressure** The pressure exerted by the atmosphere on the surface of the Earth due to the weight of the air.
- **Boiling point** The temperature at which the pressure of the gas created above a liquid equals atmospheric pressure.
- **Condensation** The change of a vapour or a gas into a liquid. This process is accompanied by the evolution of heat.
- **Diffusion** The process by which different substances mix as a result of the random motions of their particles.

- **Evaporation** A process occurring at the surface of a liquid involving the change of state of a liquid into a vapour at a temperature below the boiling point.
- **Kinetic theory** A theory which accounts for the bulk properties of matter in terms of the constituent particles.
- **Matter** Anything which occupies space and has a mass.
- **Melting point** The temperature at which a solid begins to liquefy. Pure substances have a sharp melting point.
- **Solids, liquids and gases** The three states of matter to which all substances belong.
- **Sublimation** The direct change of state from solid to gas and the reverse process.



# The particulate nature of matter

## Additional questions

- 1 a Draw diagrams to show the arrangement of particles in:
- (i) solid lead
  - (ii) molten lead
  - (iii) gaseous lead.
- b Explain how the particles move in these three states of matter.
- c Explain, using the kinetic theory, what happens to the particles in oxygen as it is cooled down.

2 Explain the meaning of each of the following terms. In your answer include an example to help with your explanation.

- |                    |                  |
|--------------------|------------------|
| a Expansion.       | d Sublimation.   |
| b Contraction.     | e Diffusion.     |
| c Physical change. | f Random motion. |

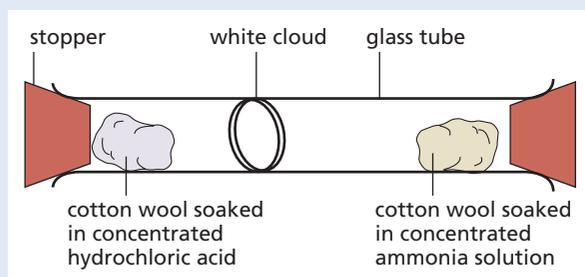
3 a Why do solids not diffuse?

b Give two examples of diffusion of gases and liquids found in the house.

4 Use the kinetic theory to explain the following:

- a When you take a block of butter out of the fridge, it is quite hard. However, after 15 minutes it is soft enough to spread.
- b When you come home from school and open the door you can smell your tea being cooked.
- c A football is blown up until it is hard on a hot summer's day. In the evening the football feels softer.
- d When a person wearing perfume enters a room it takes several minutes for the smell to reach the back of the room.
- e A windy day is a good drying day.

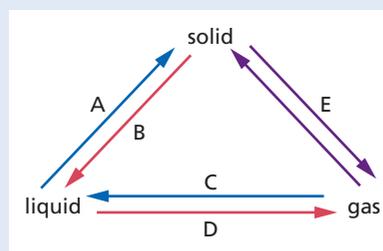
5 The apparatus shown below was set up.



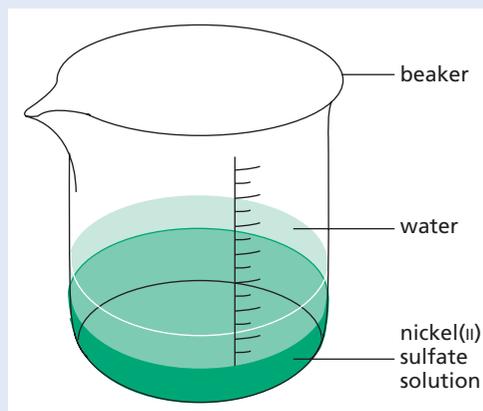
Give explanations for the following observations.

- a The formation of a white cloud.
- b It took a few minutes before the white cloud formed.

- c The white cloud formed further from the cotton wool soaked in ammonia.
  - d Cooling the concentrated ammonia and hydrochloric acid before carrying out the experiment increased the time taken for the white cloud to form.
- 6 The following diagram shows the three states of matter and how they can be interchanged.



- a Name the changes A to E.
  - b Name a substance which will undergo change E.
  - c Name a substance which will undergo changes from solid to liquid to gas between 0°C and 100°C.
  - d Describe what happens to the particles of the solid during change E.
  - e Which of the changes A to E will involve:
    - (i) an input of heat energy?
    - (ii) an output of heat energy?
- 7 Some nickel(II) sulfate solution was carefully placed in the bottom of a beaker of water. The beaker was then covered and left for several days.



- a Describe what you would see after:
  - (i) a few hours
  - (ii) several days.
- b Explain your answer to a using your ideas of the kinetic theory of particles.
- c What is the name of the physical process that takes place in this experiment?

# 2

## Elements, compounds and experimental techniques

### Elements

Atoms – the smallest particles  
Molecules

### Compounds

More about formulae  
Balancing chemical equations  
Instrumental techniques

### Mixtures

What is the difference between mixtures and compounds?

### Separating mixtures

Separating solid/liquid mixtures  
Separating liquid/liquid mixtures

Separating solid/solid mixtures  
Criteria for purity

### Accuracy in experimental work in the laboratory

Apparatus used for measurement in chemistry

### Gels, sols, foams and emulsions

### Mixtures for strength

Composite materials

### Checklist

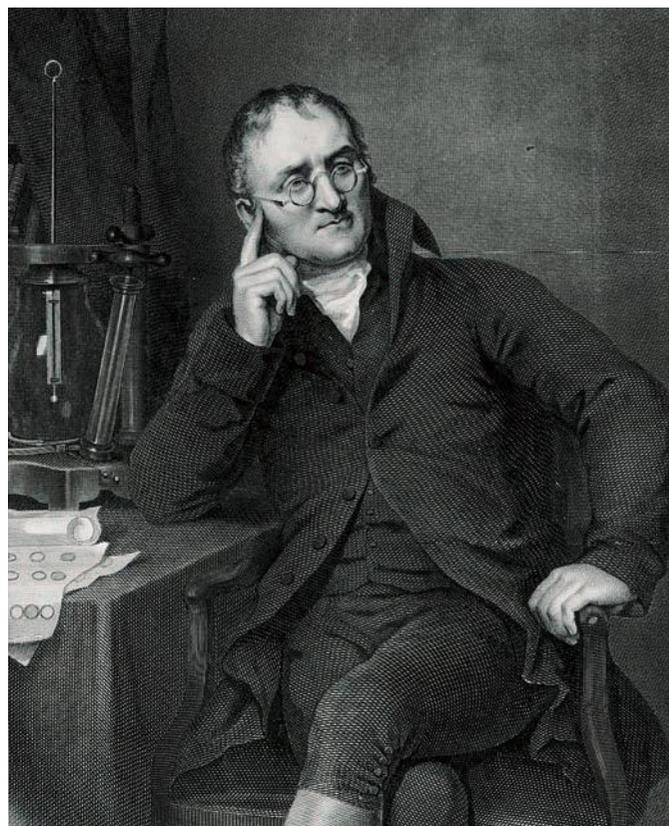
### Additional questions

The universe is made up of a very large number of substances (Figure 2.1), and our own world is no exception. If this vast array of substances is examined more closely, it is found that they are made up of some basic substances which were given the name **elements** in 1661 by Robert Boyle.



**Figure 2.1** The planets in the universe are made of millions of substances. These are made up mainly from just 91 elements which occur naturally on the Earth.

In 1803, John Dalton (Figure 2.2) suggested that each element was composed of its own kind of particles, which he called **atoms**. Atoms are much too small to be seen. We now know that about  $20 \times 10^6$  of them would stretch over a length of only 1 cm.



**Figure 2.2** John Dalton (1766–1844).

## ● Elements

Robert Boyle used the name element for any substance that cannot be broken down further, into a simpler substance. This definition can be extended to include the fact that each element is made up of only one kind of atom. The word atom comes from the Greek word *atomos* meaning ‘unsplittable’.

For example, aluminium is an element which is made up of only aluminium atoms. It is not possible to obtain a simpler substance chemically from the aluminium atoms. You can only make more complicated substances from it, such as aluminium oxide, aluminium nitrate or aluminium sulfate.

There are 118 elements which have now been identified. Twenty-seven of these do not occur in nature and have been made artificially by scientists. They include elements such as curium and unnilpentium. Ninety-one of the elements occur naturally and range from some very reactive gases, such as fluorine and chlorine, to gold and platinum, which are unreactive elements.

All elements can be classified according to their various properties. A simple way to do this is to classify them as **metals** or **non-metals** (Figures 2.3 and 2.4, p. 12). Table 2.1 shows the physical data for some common metallic and non-metallic elements.

You will notice that many metals have high densities, high melting points and high boiling points, and that most non-metals have low densities, low melting points and low boiling points. Table 2.2 summarises the different properties of metals and non-metals.

A discussion of the chemical properties of metals is given in Chapters 9 and 10. The chemical properties of certain non-metals are discussed in Chapters 9, 12 and 13.

**Table 2.1** Physical data for some metallic and non-metallic elements at room temperature and pressure.

Element	Metal or non-metal	Density/ g cm <sup>-3</sup>	Melting point/°C	Boiling point/°C
Aluminium	Metal	2.70	660	2580
Copper	Metal	8.92	1083	2567
Gold	Metal	19.29	1065	2807
Iron	Metal	7.87	1535	2750
Lead	Metal	11.34	328	1740
Magnesium	Metal	1.74	649	1107
Nickel	Metal	8.90	1453	2732
Silver	Metal	10.50	962	2212
Zinc	Metal	7.14	420	907
Carbon	Non-metal	2.25	Sublimes at 3642	
Hydrogen	Non-metal	0.07 <sup>a</sup>	-259	-253
Nitrogen	Non-metal	0.88 <sup>b</sup>	-210	-196
Oxygen	Non-metal	1.15 <sup>c</sup>	-218	-183
Sulfur	Non-metal	2.07	113	445

Source: Earl B., Wilford L.D.R. Chemistry data book. Nelson Blackie, 1991 <sup>a</sup> At -254°C <sup>b</sup> At -197°C <sup>c</sup> At -184°C.



**a** Gold is very decorative.



**b** Aluminium has many uses in the aerospace industry.



**c** These coins contain nickel.

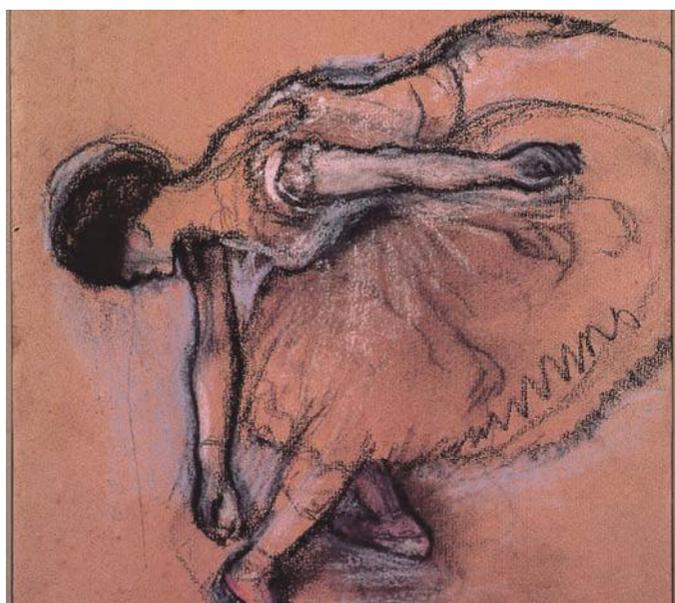
**Figure 2.3** Some metals.

**Table 2.2** How the properties of metals and non-metals compare.

Property	Metal	Non-metal
Physical state at room temperature	Usually solid (occasionally liquid)	Solid, liquid or gas
Malleability	Good	Poor – usually soft or brittle
Ductility	Good	
Appearance (solids)	Shiny (lustrous)	Dull
Melting point	Usually high	Usually low
Boiling point	Usually high	Usually low
Density	Usually high	Usually low
Conductivity (thermal and electrical)	Good	Very poor



a A premature baby needs oxygen.



b Artists often use charcoal (carbon) to produce an initial sketch.



c Neon is used in advertising signs

Figure 2.4 Some non-metals.

## Atoms – the smallest particles

Everything is made up of billions of atoms. The atoms of all elements are extremely small; in fact they are too small to be seen. The smallest atom known is hydrogen, with each atom being represented as a sphere having a diameter of 0.000 000 07 mm (or  $7 \times 10^{-8}$  mm) (Table 2.3). Atoms of different elements have different diameters as well as different masses. How many atoms of hydrogen would have to be placed side by side along the edge of your ruler to fill just one of the 1 mm divisions?

Table 2.3 Sizes of atoms.

Atom	Diameter of atom/mm
Hydrogen	$7 \times 10^{-8}$
Oxygen	$12 \times 10^{-8}$
Sulfur	$20.8 \times 10^{-8}$

Chemists use shorthand symbols to label the elements and their atoms. The symbol consists of one, two or three letters, the first of which must be a capital. Where several elements have the same initial letter, a second letter of the name or subsequent letter is added. For example, **C** is used for **carbon**, **Ca** for **calcium** and **Cl** for **chlorine**. Some symbols seem to have no relationship to the name of the element, for example **Na** for **sodium** and **Pb** for **lead**. These symbols come from their Latin names, natrium for sodium and plumbum for lead. A list of some common elements and their symbols is given in Table 2.4.

## Molecules

The atoms of some elements are joined together in small groups. These small groups of atoms are called **molecules**. For example, the atoms of the elements hydrogen, oxygen, nitrogen, fluorine, chlorine, bromine and iodine are each joined in pairs and they are known as **diatomic** molecules. In the case of phosphorus and sulfur the atoms are joined in larger numbers, four and eight respectively ( $P_4$ ,  $S_8$ ). In chemical shorthand the molecule of chlorine shown in Figure 2.5 is written as  $Cl_2$ .

**Table 2.4** Some common elements and their symbols. The Latin names of some of the elements are given in brackets.

Element	Symbol	Physical state at room temperature and pressure
Aluminium	Al	Solid
Argon	Ar	Gas
Barium	Ba	Solid
Boron	B	Solid
Bromine	Br	Liquid
Calcium	Ca	Solid
Carbon	C	Solid
Chlorine	Cl	Gas
Chromium	Cr	Solid
Copper (Cuprum)	Cu	Solid
Fluorine	F	Gas
Germanium	Ge	Solid
Gold (Aurum)	Au	Solid
Helium	He	Gas
Hydrogen	H	Gas
Iodine	I	Solid
Iron (Ferrum)	Fe	Solid
Lead (Plumbum)	Pb	Solid
Magnesium	Mg	Solid
Mercury (Hydragyrum)	Hg	Liquid
Neon	Ne	Gas
Nitrogen	N	Gas
Oxygen	O	Gas
Phosphorus	P	Solid
Potassium (Kalium)	K	Solid
Silicon	Si	Solid
Silver (Argentum)	Ag	Solid
Sodium (Natrium)	Na	Solid
Sulfur	S	Solid
Tin (Stannum)	Sn	Solid
Zinc	Zn	Solid

The complete list of the elements with their corresponding symbols is shown in the Periodic Table on p. 294.

The gaseous elements helium, neon, argon, krypton, xenon and radon are composed of separate and individual atoms. When an element exists as separate atoms, then the molecules are said to be **monatomic**. In chemical shorthand these monatomic molecules are written as He, Ne, Ar, Kr, Xe and Rn respectively.



a As a letter-and-stick model.



b As a space-filling model.

**Figure 2.5** A chlorine molecule.

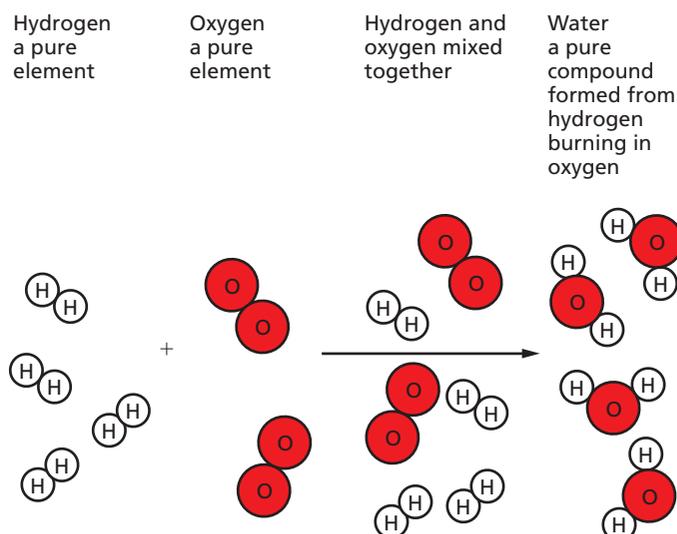
Molecules are not always formed by atoms of the same type joining together. For example, water exists as molecules containing oxygen and hydrogen atoms.

### Questions

- How would you use a similar chemical shorthand to write a representation of the molecules of iodine and fluorine?
- Using the Periodic Table on p. 294 write down the symbols for the following elements and give their physical states at room temperature:  
a chromium    b krypton    c osmium.

## Compounds

Compounds are pure substances which are formed when two or more elements chemically combine together. Water is a simple compound formed from the elements hydrogen and oxygen (Figure 2.6). This combining of the elements can be represented by a word equation:



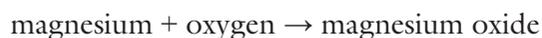
**Figure 2.6** The element hydrogen reacts with the element oxygen to produce the compound water.

Water molecules contain two atoms of hydrogen and one atom of oxygen, and hence water has the **chemical formula**  $\text{H}_2\text{O}$ . Elements other than hydrogen will also react with oxygen to form compounds called oxides. For example, magnesium reacts violently with oxygen gas to form the white powder magnesium oxide (Figure 2.7). This reaction is accompanied by a release of energy as new chemical bonds are formed.



**Figure 2.7** Magnesium burns brightly in oxygen to produce magnesium oxide.

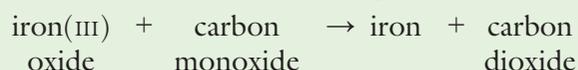
When a new substance is formed during a chemical reaction, a **chemical change** has taken place.



When substances such as hydrogen and magnesium combine with oxygen in this way they are said to have been **oxidised**. The process is known as **oxidation**.

**Reduction** is the opposite of oxidation. In this process oxygen is removed instead of being added.

A **redox** reaction is one which involves the two processes of reduction and oxidation. For example, the oxygen has to be removed in the extraction of iron from iron(III) oxide. This can be done in a blast furnace with carbon monoxide. The iron(III) oxide loses oxygen to the carbon monoxide and is reduced to iron. Carbon monoxide is the **reducing agent**. A reducing agent is a substance that reduces another substance during a redox reaction. Carbon monoxide is oxidised to carbon dioxide by the iron(III) oxide. The iron(III) oxide is the **oxidising agent**. An oxidising agent is a substance which oxidises another substance during a redox reaction.



For a further discussion of oxidation and reduction see Chapter 3 (p. 39) and Chapter 5 (p. 73).

Both **reduction** and **oxidation** have taken place in this chemical process, and so this is known as a **redox** reaction.

### More about formulae

The formula of a compound is made up from the symbols of the elements present and numbers to show the ratio in which the different atoms are present. Carbon dioxide has the formula  $\text{CO}_2$ . This tells you that it contains one carbon atom for every two oxygen atoms. The 2 in the formula tells you that there are two oxygen atoms present in each molecule of carbon dioxide. For further discussion see p. 43.

Table 2.5 shows the names and formulae of some common compounds which you will meet in your study of chemistry.

**Table 2.5** Names and formulae of some common compounds.

Compound	Formula
Ammonia	$\text{NH}_3$
Calcium hydroxide	$\text{Ca(OH)}_2$
Carbon dioxide	$\text{CO}_2$
Copper sulfate	$\text{CuSO}_4$
Ethanol (alcohol)	$\text{C}_2\text{H}_5\text{OH}$
Glucose	$\text{C}_6\text{H}_{12}\text{O}_6$
Hydrochloric acid	$\text{HCl}$
Nitric acid	$\text{HNO}_3$
Sodium carbonate	$\text{Na}_2\text{CO}_3$
Sodium hydroxide	$\text{NaOH}$
Sulfuric acid	$\text{H}_2\text{SO}_4$

The ratio of atoms within a chemical compound is usually constant. Compounds are made up of fixed proportions of elements: they have a fixed composition. Chemists call this the **Law of constant composition**.

## Balancing chemical equations

Word equations are a useful way of representing chemical reactions but a better and more useful method is to produce a **balanced chemical equation**. This type of equation gives the formulae of the reactants and the products as well as showing the relative numbers of each particle involved.

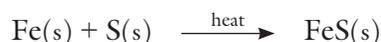
Balanced equations often include the physical state symbols:

(s) = solid, (l) = liquid, (g) = gas, (aq) = aqueous solution

The word equation to represent the reaction between iron and sulfur is:

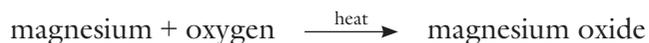


When we replace the words with symbols for the reactants and the products and include their physical state symbols, we obtain:



Since there is the same number of each type of atom on both sides of the equation this is a **balanced** chemical equation.

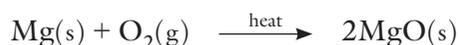
In the case of magnesium reacting with oxygen, the word equation was:



When we replace the words with symbols for the reactants and the products and include their physical state symbols, it is important to remember that oxygen is a diatomic molecule:



In the equation there are two oxygen atoms on the left-hand side ( $\text{O}_2$ ) but only one on the right ( $\text{MgO}$ ). We cannot change the formula of magnesium oxide, so to produce the necessary two oxygen atoms on the right-hand side we will need  $2\text{MgO}$  – this means  $2 \times \text{MgO}$ . The equation now becomes:



There are now two atoms of magnesium on the right-hand side and only one on the left. By placing a 2 in front of the magnesium, we obtain the following balanced chemical equation:



This balanced chemical equation now shows us that two atoms of magnesium react with one molecule of oxygen gas when heated to produce two units of magnesium oxide.

## Instrumental techniques

Elements and compounds can be detected and identified by a variety of instrumental methods. Scientists have developed instrumental techniques that allow us to probe and discover which elements are present in the substance as well as how the atoms are arranged within the substance.

Many of the instrumental methods that have been developed are quite sophisticated. Some methods are suited to identifying elements. For example, atomic absorption spectroscopy allows the element to be identified and also allows the quantity of the element that is present to be found (Figure 2.8).



**Figure 2.8** This instrument allows the quantity of a particular element to be found. It is used extensively throughout industry for this purpose. It will allow even tiny amounts of a particular element to be found.

Some methods are particularly suited to the identification of compounds. For example, infrared spectroscopy is used to identify compounds by showing the presence of particular groupings of atoms (Figure 2.9).



**Figure 2.9** This is a modern infrared spectrometer. It is used in analysis to obtain a so-called fingerprint spectrum of a substance that will allow the substance to be identified.

Infrared spectroscopy is used by the pharmaceutical industry to identify and discriminate between drugs that are similar in structure, for example penicillin-type drugs. Used both with organic and inorganic molecules, this method assumes that each compound has a unique infrared spectrum. Samples can be solid, liquid or gas and are usually tiny. However, Ne, He, O<sub>2</sub>, N<sub>2</sub> or H<sub>2</sub> cannot be used.

This method is also used to monitor environmental pollution, and has biological uses in monitoring tissue physiology including oxygenation, respiratory status and blood flow damage.

Forensic scientists make use of both these techniques because they are very accurate but they only require tiny amounts of sample – often only small amounts of sample are found at crime scenes. Other techniques utilised are nuclear magnetic resonance spectroscopy and ultraviolet/visible spectroscopy.

### Questions

- Write the word and balanced chemical equations for the reactions which take place between:
  - calcium and oxygen
  - copper and oxygen.
- Write down the ratio of the atoms present in the formula for each of the compounds shown in Table 2.5.
- Iron is extracted from iron(III) oxide in a blast furnace by a redox reaction. What does the term 'redox reaction' mean?
- Identify the oxidising and reducing agents in the following reactions:
  - copper(II) oxide + hydrogen → copper + water
  - tin(II) oxide + carbon → tin + carbon dioxide
  - $\text{PbO(s)} + \text{H}_2\text{(g)} \rightarrow \text{Pb(s)} + \text{H}_2\text{O(l)}$

## Mixtures

Many everyday things are not pure substances, they are mixtures. A mixture contains more than one substance (elements and/or compounds). An example of a common mixture is sea water (Figure 2.10).



**Figure 2.10** Sea water is a common mixture.

Other mixtures include the air, which is a mixture of elements such as oxygen, nitrogen and neon and compounds such as carbon dioxide (see Chapter 11, p. 173), and alloys such as brass, which is a mixture of copper and zinc (for a further discussion of alloys see Chapter 10, p. 165).

### What is the difference between mixtures and compounds?

There are differences between compounds and mixtures. This can be shown by considering the reaction between iron filings and sulfur. A mixture of iron filings and sulfur looks different from the individual elements (Figure 2.11). This mixture has the properties of both iron and sulfur; for example, a magnet can be used to separate the iron filings from the sulfur (Figure 2.12).

Substances in a mixture have not undergone a chemical reaction and it is possible to separate them provided that there is a suitable difference in their physical properties. If the mixture of iron and sulfur is heated a chemical reaction occurs and a new substance is formed called iron(II) sulfide (Figure 2.11). The word equation for this reaction is:





**Figure 2.11** The elements sulfur and iron at the top of the photograph, and (below) black iron(II) sulfide on the left and a mixture of the two elements on the right.



**Figure 2.12** A magnet will separate the iron from the mixture.

During the reaction heat energy is given out as new chemical bonds are formed. This is called an **exothermic** reaction and accompanies a chemical change (Chapter 6, pp. 92 and 95). The iron(II) sulfide formed has totally different properties to the mixture of iron and sulfur (Table 2.6). Iron(II) sulfide, for example, would not be attracted towards a magnet.

**Table 2.6** Different properties of iron, sulfur, an iron/sulfur mixture and iron(II) sulfide.

Substance	Appearance	Effect of a magnet	Effect of dilute hydrochloric acid
Iron	Dark grey powder	Attracted to it	Very little action when cold. When warm, a gas is produced with a lot of bubbling (effervescence)
Sulfur	Yellow powder	None	No effect when hot or cold
Iron/sulfur mixture	Dirty yellow powder	Iron powder attracted to it	Iron powder reacts as above
Iron(II) sulfide	Black solid	No effect	A foul-smelling gas is produced with some effervescence

**Table 2.7** The major differences between mixtures and compounds.

Mixture	Compound
It contains two or more substances	It is a single substance
The composition can vary	The composition is always the same
No chemical change takes place when a mixture is formed	When the new substance is formed it involves chemical change
The properties are those of the individual elements/compounds	The properties are very different to those of the component elements
The components may be separated quite easily by physical means	The components can only be separated by one or more chemical reactions

In iron(II) sulfide, FeS, one atom of iron has combined with one atom of sulfur. No such ratio exists in a mixture of iron and sulfur, because the atoms have not chemically combined. Table 2.7 summarises how mixtures and compounds compare.

Some common mixtures are discussed in Chapter 10 (p. 165) and Chapter 11 (p. 173).

### Question

- 1 Make a list of some other common mixtures, stating what they are mixtures of.

## ● Separating mixtures

Many mixtures contain useful substances mixed with unwanted material. In order to obtain these useful substances, chemists often have to separate them from the impurities. Chemists have developed many different methods of separation, particularly for separating compounds from complex mixtures. Which separation method they use depends on what is in the mixture and the properties of the substances present. It also depends on whether the substances to be separated are solids, liquids or gases.

### Separating solid/liquid mixtures

If a solid substance is added to a liquid it may **dissolve** to form a **solution**. In this case the solid is said to be **soluble** and is called the **solute**. The liquid it has dissolved in is called the **solvent**. An example of this type of process is when sugar is added to tea or coffee. What other examples can you think of where this type of process takes place?

Sometimes the solid does not dissolve in the liquid. This solid is said to be **insoluble**. For example, tea leaves themselves do not dissolve in boiling water when tea is made from them, although the soluble materials from which tea is made are seen to dissolve from them.

## Filtration

When a cup of tea is poured through a tea strainer you are carrying out a **filtering** process. **Filtration** is a common separation technique used in chemistry laboratories throughout the world. It is used when a solid needs to be separated from a liquid. For example, sand can be separated from a mixture with water by filtering through filter paper as shown in Figure 2.13.



**Figure 2.13** It is important when filtering not to overfill the filter paper.

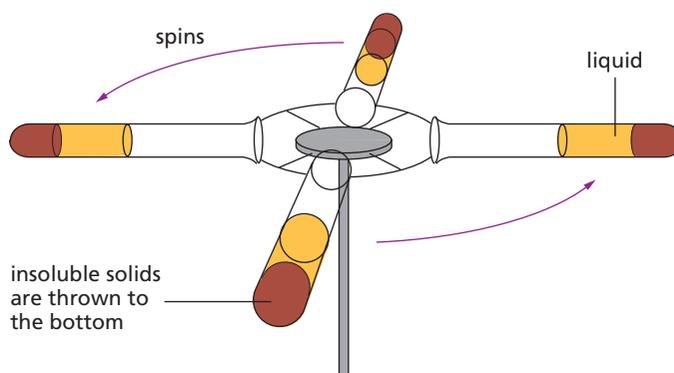
The filter paper contains holes that, although too small to be seen, are large enough to allow the molecules of water through but not the sand particles. It acts like a sieve. The sand gets trapped in the filter paper and the water passes through it. The sand is called the **residue** and the water is called the **filtrate**.

## Decanting

Vegetables do not dissolve in water. When you have boiled some vegetables it is easy to separate them from the water by pouring it off. This process is called **decanting**. This technique is used quite often to separate an insoluble solid, which has settled at the bottom of a flask, from a liquid.

## Centrifuging

Another way to separate a solid from a liquid is to use a **centrifuge**. This technique is sometimes used instead of filtration. It is usually used when the solid particles are so small that they spread out (disperse) throughout the liquid and remain in **suspension**. They do not settle to the bottom of the container, as heavier particles would do, under the force of gravity. The technique of **centrifuging** or **centrifugation** involves the suspension being spun round very fast in a centrifuge so that the solid gets flung to the bottom of the tube (Figure 2.14a and b).



**a** The sample is spun round very fast and the solid is flung to the bottom of the tube.



**b** An open centrifuge.

**Figure 2.14**

The pure liquid can be decanted after the solid has been forced to the bottom of the tube. This method of separation is used extensively to separate blood cells from blood plasma (Figure 2.15). In this case, the solid particles (the blood cells) are flung to the bottom of the tube, allowing the liquid plasma to be decanted.



**Figure 2.15** Whole blood (top) is separated by centrifuging into blood cells and plasma (bottom).

## Evaporation

If the solid has dissolved in the liquid it cannot be separated by filtering or centrifuging. Instead, the solution can be heated so that the liquid evaporates completely and leaves the solid behind. The simplest way to obtain salt from its solution is by slow evaporation as shown in Figure 2.16.



**Figure 2.16** Apparatus used to slowly evaporate a solvent.

## Crystallisation

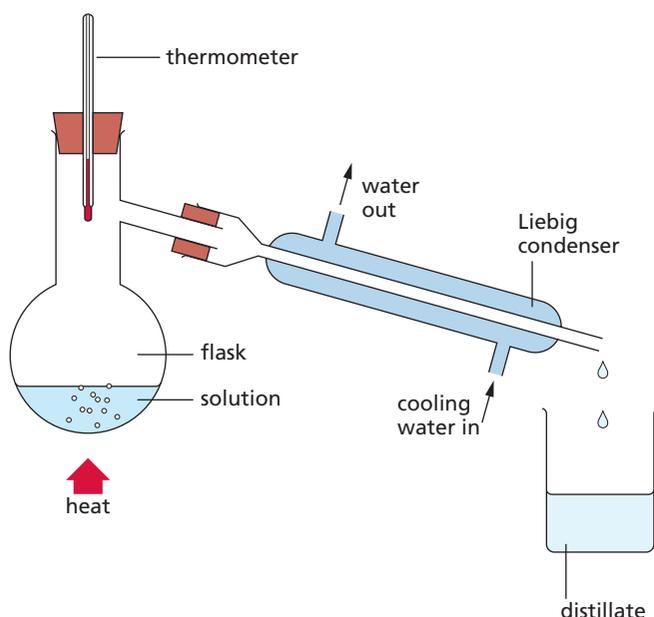
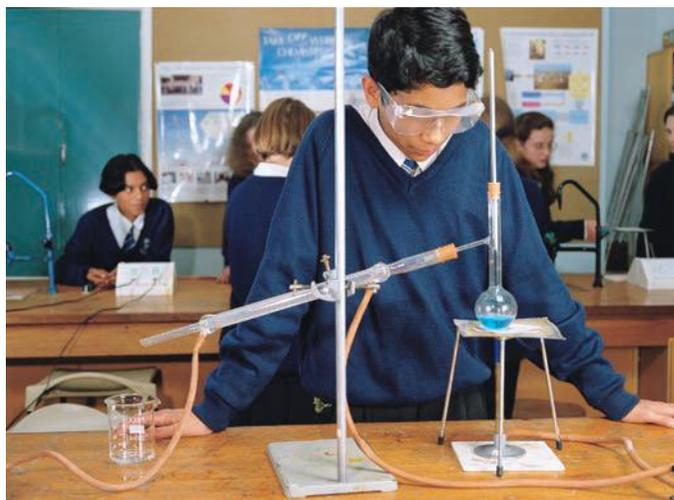
In many parts of the world salt is obtained from sea water on a vast scale. This is done by using the heat of the sun to evaporate the water to leave a saturated solution of salt known as brine. A **saturated solution** is defined as one that contains as much solute as can be dissolved at a particular temperature. When the solution is saturated the salt begins to **crystallise**, and it is removed using large scoops (Figure 2.17).



**Figure 2.17** Salt is obtained in north-eastern Brazil by evaporation of sea water.

## Simple distillation

If we want to obtain the solvent from a solution, then the process of **distillation** can be carried out. The apparatus used in this process is shown in Figure 2.18.



**Figure 2.18** Water can be obtained from salt water by distillation.

Water can be obtained from salt water using this method. The solution is heated in the flask until it boils. The steam rises into the Liebig condenser, where it condenses back into water. The salt is left behind in the flask. In hot and arid countries such as Saudi Arabia this sort of technique is used on a much larger scale to obtain pure water for drinking (Figure 2.19). This process is carried out in a desalination plant.



**Figure 2.19** This plant produces large quantities of drinking water in Saudi Arabia.

## Separating liquid/liquid mixtures

In recent years there have been many oil tanker disasters, just like the one shown in Figure 2.20. These have resulted in millions of litres of oil being washed into the sea. Oil and water do not mix easily. They are said to be **immiscible**. When cleaning up disasters of this type, a range of chemicals can be added to the oil to make it more soluble. This results in the oil and water mixing with each other. They are now said to be **miscible**. The following techniques can be used to separate mixtures of liquids.



**Figure 2.20** Millions of litres of oil are spilt in tanker disasters and cleaning up is a slow and costly process.

## Liquids which are immiscible

If two liquids are immiscible they can be separated using a **separating funnel**. The mixture is poured into the funnel and the layers allowed to separate.

The lower layer can then be run off by opening the tap as shown in Figure 2.21.



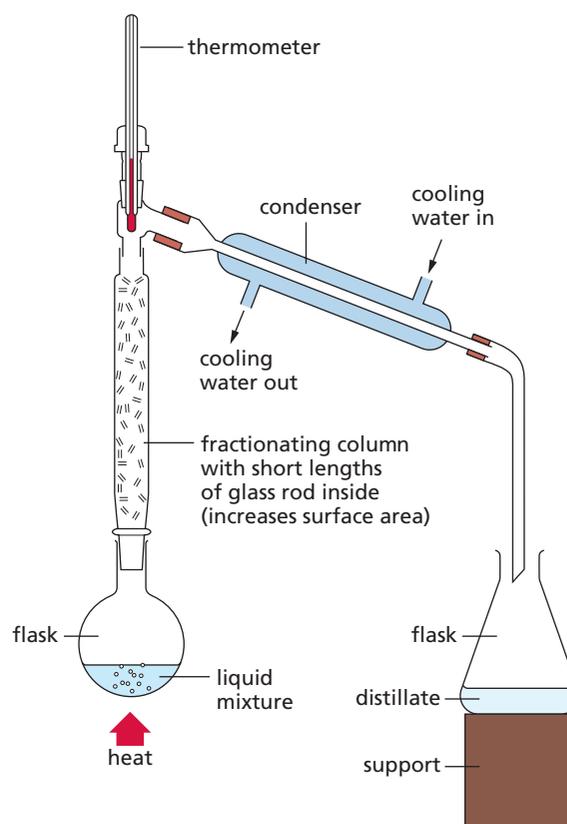
**Figure 2.21** The pink liquid is more dense than the clear oil and so sinks to the bottom of the separating funnel. When the tap is opened the pink liquid can be run off.

### Liquids which are miscible

If miscible liquids are to be separated, then this can be done by **fractional distillation**. The apparatus used for this process is shown in the photo and diagram in Figure 2.22, and could be used to separate a mixture of ethanol and water.

Fractional distillation relies upon the liquids having different boiling points. When an ethanol and water mixture is heated the vapours of ethanol and water boil off at different temperatures and can be condensed and collected separately.

Ethanol boils at  $78^{\circ}\text{C}$  whereas water boils at  $100^{\circ}\text{C}$ . When the mixture is heated the vapour produced is mainly ethanol with some steam. Because water has the higher boiling point of the two, it condenses out from the mixture with ethanol. This is what takes place in the fractionating column. The water condenses and drips back into the flask while the ethanol vapour moves up the column and into the condenser, where it condenses into liquid ethanol and is collected in the receiving flask as the distillate. When all the ethanol has distilled over, the temperature reading on the thermometer rises



**Figure 2.22** Typical fractional distillation apparatus.

steadily to  $100^{\circ}\text{C}$ , showing that the steam is now entering the condenser. At this point the receiver can be changed and the condensing water can now be collected.

Fractional distillation is used to separate miscible liquids such as those in crude oil (see Figure 2.23a and p. 90), and the technique can also separate individual gases, such as nitrogen, from the mixture we call air (see Figure 2.23b and p. 174).



a Fractional distillation unit for crude oil.



b Gases from the air are extracted in this fractional distillation plant.

Figure 2.23

## Separating solid/solid mixtures

You saw earlier in this chapter (p. 16) that it was possible to separate iron from sulfur using a magnet. In that case we were using one of the physical properties of iron, that is, the fact that it is magnetic. In a similar way, it is possible to separate scrap iron from other metals by using a large electromagnet like the one shown in Figure 2.24.



Figure 2.24 Magnetic separation of iron-containing materials.

It is essential that when separating solid/solid mixtures you pay particular attention to the individual physical properties of the components. If, for example, you wish to separate two solids, one of which sublimes, then this property should dictate the method you employ.

In the case of an iodine/salt mixture the iodine sublimes but salt does not. Iodine can be separated by heating the mixture in a fume cupboard as shown in Figure 2.25. The iodine sublimes and re-forms on the cool inverted funnel.

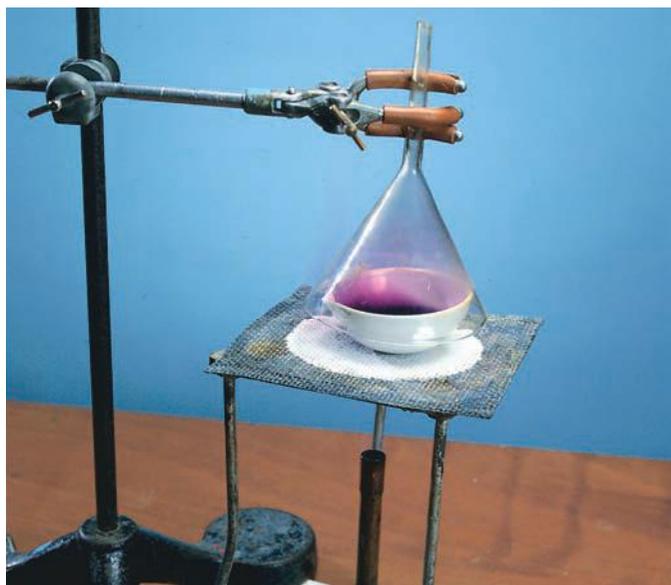


Figure 2.25 Apparatus used to separate an iodine/salt mixture. The iodine sublimes on heating.

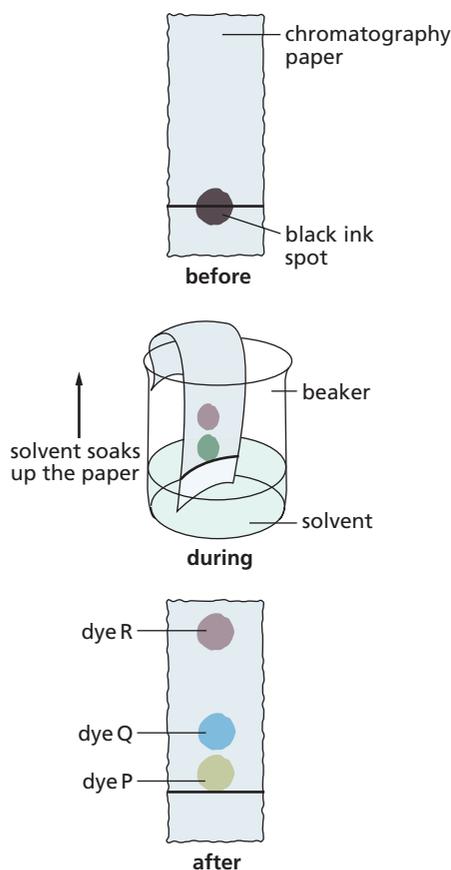
## Chromatography

What happens if you have to separate two or more solids that are soluble? This type of problem is encountered when you have mixtures of coloured materials such as inks and dyes. A technique called **chromatography** is widely used to separate these materials so that they can be identified.

There are several types of chromatography; however, they all follow the same basic principles. The simplest kind is paper chromatography. To separate the different coloured dyes in a sample of black ink, a spot of the ink is put on to a piece of chromatography paper. This paper is then set in a suitable solvent as shown in Figure 2.26.



a Chromatographic separation of black ink.



b The black ink separates into three dyes: P, Q and R.

Figure 2.26

As the solvent moves up the paper, the dyes are carried with it and begin to separate. They separate because the substances have different solubilities in the solvent and are absorbed to different degrees by the chromatography paper. As a result, they are separated gradually as the solvent moves up the paper. The **chromatogram** in Figure 2.26b shows how the ink contains three dyes, P, Q and R.

Numerical measurements (retardation factors) known as  **$R_f$  values** can be obtained from chromatograms. An  $R_f$  value is defined as the ratio of the distance travelled by the solute (for example P, Q or R) to the distance travelled by the solvent.

Chromatography and electrophoresis (separation according to electrical charge) are used extensively in medical research and forensic science laboratories to separate a variety of mixtures (Figure 2.27).

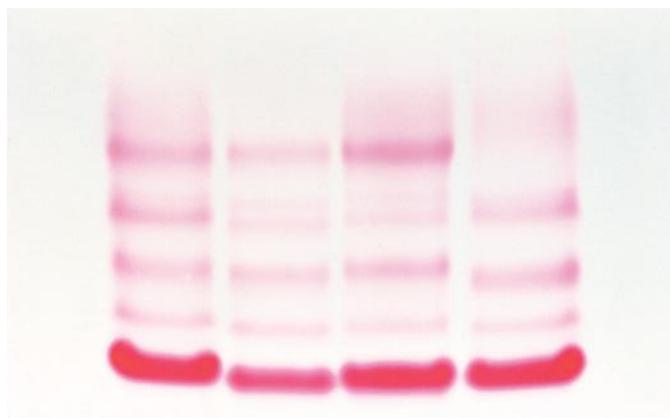


Figure 2.27 Protein samples are separated by electrophoresis in medical research laboratories.

The substances to be separated do not have to be coloured. Colourless substances can be made visible by spraying the chromatogram with a **locating agent**. The locating agent will react with the colourless substances to form a coloured product. In other situations the position of the substances on the chromatogram may be located using ultraviolet light.

## Solvent extraction

Sugar can be obtained from crushed sugar cane by adding water. The water dissolves the sugar from the sugar cane (Figure 2.28). This is an example

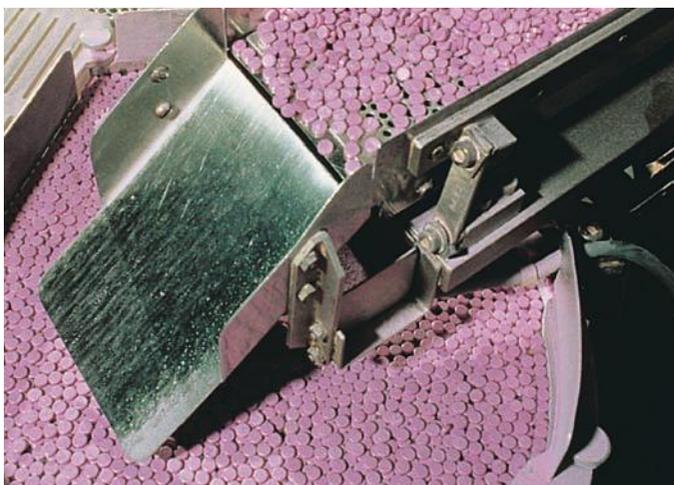
of **solvent extraction**. In a similar way some of the green substances can be removed from ground-up grass using ethanol. The substances are extracted from a mixture by using a solvent which dissolves only those substances required.



**Figure 2.28** Cutting sugar cane, from which sugar can be extracted by using a suitable solvent.

## Criteria for purity

Drugs are manufactured to a very high degree of purity (Figure 2.29). To ensure that the highest possible purity is obtained, the drugs are dissolved in a suitable solvent and subjected to fractional crystallisation.



**Figure 2.29** Drugs are manufactured to a high degree of purity by fractional crystallisation.

It is illegal to put anything harmful into food. Also, government legislation requires that a lot of testing takes place before a new pharmaceutical is marketed.

Throughout the chemical, pharmaceutical and food industries it is essential that the substances used are pure. The purity of a substance can be gauged by:

- its melting point – if it is a pure solid it will have a sharp melting point. If an impurity is present then melting takes place over a range of temperatures.
- its boiling point – if it is a pure liquid the temperature will remain steady at its boiling point. If the substance is impure then the mixture will boil over a temperature range.
- chromatography – if it is a pure substance it will produce only one well-defined spot on a chromatogram. If impurities are present then several spots will be seen on the chromatogram (see Figure 2.26, p. 23).



**Figure 2.30** These pharmaceuticals must have been through a lot of testing before they can be sold in a chemist's shop.

## Questions

- 1 Use your research techniques (including the Internet) to obtain as many examples as you can in which a centrifuge is used.
- 2 What is the difference between simple distillation and fractional distillation?
- 3 Describe how you would use chromatography to show whether blue ink contains a single pure dye or a mixture of dyes.
- 4 Explain the following terms, with the aid of examples:
 

a miscible	b immiscible
c evaporation	d condensation
e solvent extraction.	
- 5 Devise a method for obtaining salt (sodium chloride) from sea water in the school laboratory.
- 6 What criteria can be used to test the purity of a substance?

## ● Accuracy in experimental work in the laboratory

Scientists find out about the nature of materials by carrying out experiments in a laboratory. Many of these experiments require apparatus that you have used in your study of chemistry to date. Certainly a knowledge and understanding of the use of this scientific apparatus is required for successful experimentation and investigations that you may carry out in your further study of chemistry. Much of the work involves **accurate** measurements with particular pieces of apparatus in particular experiments, many of which are shown in the section below.

### Apparatus used for measurement in chemistry

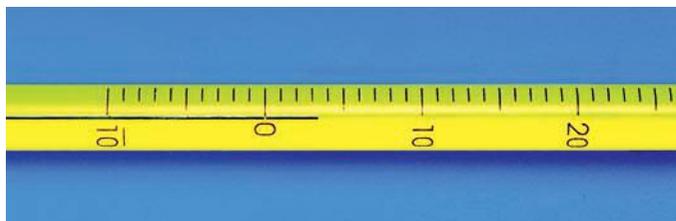
#### Measurement of time



**Figure 2.31** This stopwatch can be used to measure the time passed in a chemical reaction.

Experiments involving rates of reaction will require the use of an accurate stopwatch – one that measures to a hundredth of a second. The units of time are hours (h), minutes (min) and seconds (s).

#### Measurement of temperature



**Figure 2.32** A thermometer can be used to measure temperature.

The most commonly used thermometers in a laboratory are alcohol-in-glass. However, mercury in-glass thermometers can be used but should be handled with great care. The mercury inside them is poisonous and should not be handled if a thermometer breaks. The units of temperature are those of the Celsius scale. This scale is based on the temperature at which water freezes and boils, that is:

the freezing point of water is  $0^{\circ}\text{C}$  whilst the boiling point of water is  $100^{\circ}\text{C}$ .

For accuracy the thermometer should be capable of being read to a tenth of a degree Celsius. The usual thermometer used is that shown in the photograph that measures accurately between  $-10^{\circ}$  and  $110^{\circ}\text{C}$ . When reading the thermometer always ensure that your eye is at the same level as the liquid meniscus in the thermometer to ensure there are no parallax effects. The meniscus is the way that the liquid curves at the edges of the capillary in which the liquid is held in the thermometer.

#### Measurement of mass



**Figure 2.33** An electronic balance can be used to measure the mass of reagents.

There are many different electronic balances which can be used. The important detail with any of them is that they are accurate to one hundredth of a gram.

The units for measuring mass are grams (g) and kilograms (kg).

$$1 \text{ kg} = 1000 \text{ g}$$

When using an electronic balance you should wait until the reading is steady before taking it.

### Measurement of volume



**Figure 2.34** The apparatus shown in the photograph is generally used in different experiments to measure volume accurately.

Different experiments involving liquids will require one or other or all the various measuring apparatus for volume. The volume of a liquid is a measure of the amount of space that it takes up. The units of volume are litres (l) and cubic centimetres (cm<sup>3</sup>).

$$1 \text{ litre} = 1000 \text{ cm}^3$$

However, some of the manufacturers of apparatus used for measuring volume use millilitres (ml). This is not a problem, however, since 1 cm<sup>3</sup> = 1 ml.

When reading the volume using one of the pieces of apparatus it is important to ensure that the apparatus is vertical and that your eye is level with the top of the meniscus of the liquid being measured.

## ● Gels, sols, foams and emulsions

Gels, sols, foams and emulsions are all examples of mixtures which are formed by mixing two substances (or phases) which cannot mix. These mixtures are often referred to as **colloids**. Colloids are formed if the suspended particles are between 1 nm and 1000 nm in size (1 nm = 1 × 10<sup>-9</sup> m).

Generally colloids cannot be separated by filtration since the size of the dispersed particles is smaller than that of the pores found in the filter paper. Look closely at the substances shown in Figure 2.35 to see examples of these mixtures.



**a** These jelly-like mixtures of solid and liquid in fruit jelly and cold custard are examples of 'gels'.



**b** Emulsion paint is an example of a 'sol'.



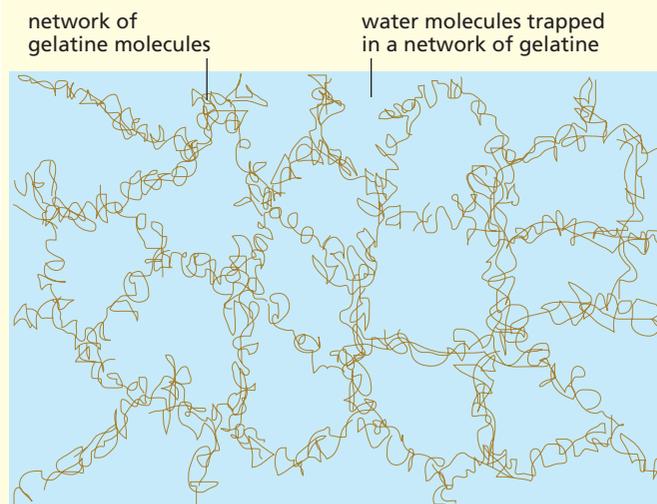
**c** These foams have been formed by trapping bubbles of gas in liquids or solids.



**d** Emulsions are formed by mixing immiscible liquids.

**Figure 2.35**

When you mix a solid with a liquid you sometimes get a gel. A gel is a semi-solid which can move around but not as freely as a liquid. Within a gel the solid makes a kind of network which traps the liquid and makes it unable to flow freely (Figure 2.36).



**Figure 2.36** The network within a gel.

A gelatine gel is made with warm water. Gelatine is a protein. Proteins are natural polymers (Chapter 15, p. 243) and the molecules of protein are very large. The large molecules disperse in water to form a gel. As the gelatine-in-water mixture cools, the gelatine molecules are attracted to each other and form a continuous network. In this way, the jelly you eat as a pudding is formed. The kind of gel which you put into your hair is made from water and an oil (Figure 2.37).

A **sol** is similar to a gel; however, the mixture will flow, for example emulsion paint, or PVA glue.



**Figure 2.37** Hair gel is a mixture of water and an oil plus a perfume.

When you pour out a glass of fizzy drink, the frothy part at the top of the drink is a gas/liquid mixture called a **foam**. The gas, carbon dioxide, has formed tiny bubbles in the liquid but has not dissolved in it. If left to stand, foams like this one collapse as the tiny bubbles join together to form bigger bubbles which then escape. It is possible to form solid foams where the gases are trapped in a solid structure. This happens in foam rubber and bread (Figure 2.38).

**Emulsions** are mixtures of liquids which are immiscible. Earlier in this chapter you found out that when two liquids are immiscible they do not mix but form two different layers. Oil and water are like this but if you shake the mixture it becomes cloudy.

The apparent mixing that you see is due to the fact that one of the liquids has been broken into tiny droplets which float suspended in the other liquid. If the mixture of oil and water is now left to stand the two layers will re-form. To make emulsions, such as mayonnaise, an **emulsifier** is used to stop the droplets joining back together again to form a separate layer. The emulsifier used when making mayonnaise is egg yolk. In many countries of the world, if you examine the ingredients on the side of many packets found in kitchen cupboards you will find that emulsifiers have ‘E-numbers’ in the range E322 to E494. For example, ammonium phosphatide E442 is used as the emulsifier in cocoa and chocolate. Other food additives such as colourings and preservatives are also given E-numbers but in different ranges to that of the emulsifiers.

It is worth noting that gels, foams and emulsions are all examples of different kinds of solutions. In true solutions the two phases completely mix together but in these systems the two phases are separate.



**Figure 2.38** Examples of solid foams.



To produce a stable colloid, the particles dispersed must not only be of the right size (1–1000 nm) but also be prevented from joining back together (coagulating). One way of doing this is to ensure that all the particles possess the same electrical charge. This causes the particles to repel one another.

A colloidal suspension can be destroyed by bringing the dispersed particles together. This process is known as **flocculation**. A method of doing this involves adding ionic substances such as aluminium chloride or aluminium sulfate to the particular colloid. The dispersed particles interact with the added highly charged ions and form particles which are large enough either to settle out under the force of gravity or simply be filtered out. During the treatment of water, aluminium sulfate is added to water prior to filtering to remove suspended solids (Figure 2.39).



**Figure 2.39** Water is treated to remove suspended solids by the addition of aluminium sulfate.

### Questions

- Explain the following terms:
 

a colloid	c foam	e sol.
b emulsifier	d 'E' number	
- Use your research skills (including the Internet) to obtain information about as many common gels, sols, foams and emulsions as you can, other than those given in the text.

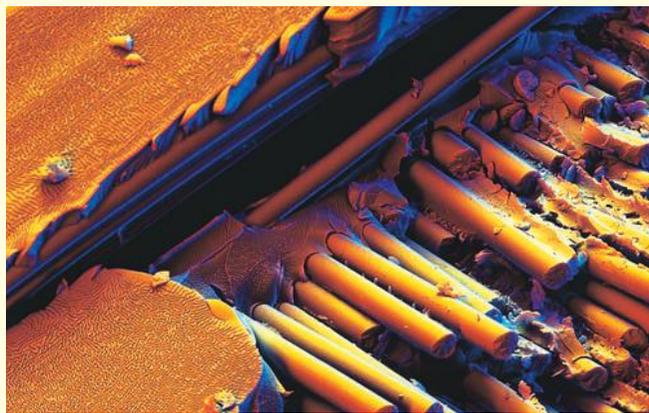
## Mixtures for strength

### Composite materials

Composite materials are those that combine the properties of two constituents in order to get the exact properties needed for a particular job.

Glass-reinforced plastic (GRP) is an example of a composite material combining the properties of two different materials. It is made by embedding short

fibres of glass in a matrix of plastic. The glass fibres give the plastic extra strength so that it does not break when it is bent or moulded into shape. The finished product has the lightness of plastic as well as the strength and flexibility of the glass fibres (Figure 2.40).

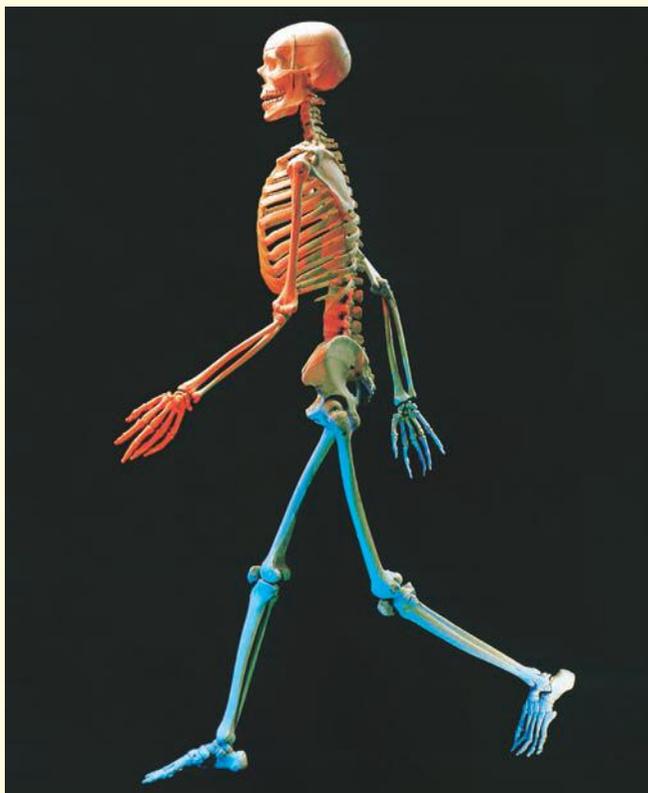


**Figure 2.40** GRP consists of glass fibres (rod shapes) embedded in plastic, in this case polyester.

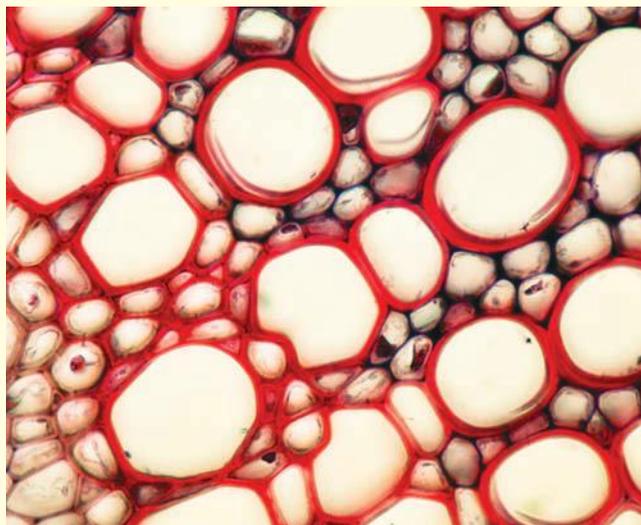


**Figure 2.41** The glass-reinforced plastic used to make boats like this is a composite material.

With a little investigation you will find that many composite materials are found in the natural world. Our bones, for example, are a composite material formed from strands of the protein collagen and the mineral calcium phosphate (Figure 2.42). The calcium phosphate is hard and therefore gives strength to the bone. Another example is wood. Wood consists of cellulose fibres mixed with lignin (Figure 2.43), which is largely responsible for the strength of the wood.



**Figure 2.42** Bone is a composite material.



**Figure 2.43** The combination of cellulose fibres and lignin makes the cell walls hard, thick and very strong. These properties reinforce the cells against collapse.

### Questions

- 1 Why are composite materials often used instead of single materials?
- 2 Using the information in the text and any other information available to you, give a use other than those already mentioned for each of the following composite materials:
  - a reinforced concrete
  - b glass-reinforced plastic
  - c laminate
  - d glass fibre.

### Checklist

After studying Chapter 2 you should know and understand the following terms.

- **Accuracy in experimental work** The accuracy of each measurement. It depends on the quality of the measuring apparatus (e.g. the thermometer or electronic balance) and on the skill of the scientists taking the measurement.
- **Atom** The smallest part of an element that can exist as a stable entity.
- **Centrifuging** The separation of the components of a mixture by rapid spinning. The denser particles are flung to the bottom of the containing tubes. The liquid can then be decanted off.
- **Chemical change** A permanent change in which a new substance is formed.
- **Chemical formula** A shorthand method of representing chemical elements and compounds.
- **Chromatography** A technique employed for the separation of mixtures of dissolved substances.

- **Colloid** System in which there are two or more phases, with one (the dispersed phase) distributed in the other (the continuous phase). One of the phases has particles in the range 1 to 1000 nm ( $1 \text{ nm} = 1 \times 10^{-9} \text{ m}$ ).
- **Composite materials** Materials which combine the properties of two substances in order to get the exact properties required for a particular job.
- **Compound** A substance formed by the combination of two or more elements in fixed proportions.
- **Crystallisation** The process of forming crystals from a liquid.
- **Decanting** The process of removing a liquid from a solid which has settled or from an immiscible heavier liquid by pouring.
- **Diatomic molecule** A molecule containing two atoms, for example hydrogen,  $\text{H}_2$ , and oxygen,  $\text{O}_2$ .
- **Distillate** The condensed vapour produced from a mixture of liquids on distillation.

- **Distillation** The process of boiling a liquid and then condensing the vapour produced back into a liquid. It is used to purify liquids and to separate mixtures of liquids.
- **Element** A substance which cannot be further divided into simpler substances by chemical methods.
- **Emulsifier** A substance used to stop the droplets that make up an emulsion joining back together again to form a separate layer.
- **Emulsion** The apparent mixing of two immiscible liquids by the use of an emulsifier which breaks down one of the liquids into tiny droplets. The droplets of this liquid float suspended in the other liquid so that they do not separate out into different layers.
- **Evaporation** When a solution is heated the solvent evaporates and leaves the solute behind.
- **Exothermic reaction** A chemical reaction in which heat energy is produced.
- **Filtrate** The liquid which passes through the filter paper during filtration.
- **Filtration** The process of separating a solid from a liquid using a fine filter paper which does not allow the solid to pass through.
- **Flocculation** The destruction of a colloidal suspension by bringing the dispersed particles together.
- **Foam** A mixture formed between a gas and a liquid. The gas forms tiny bubbles in the liquid but has not dissolved in it.
- **Fractional distillation** A distillation technique used to separate a mixture of liquids that have different boiling points.
- **Gel** A mixture formed between a solid and a liquid in which the solid forms a network which traps the liquid so that it cannot flow freely.
- **Immiscible** When two liquids form two layers when mixed together, they are said to be immiscible.
- **Insoluble** If the solute does not dissolve in the solvent it is said to be insoluble.
- **Instrumental techniques** Instrumental methods of analysis that are particularly useful when the amount of sample is very small. Examples are atomic absorption spectroscopy and infrared spectroscopy.
- **Law of constant composition** Compounds always have the same elements joined together in the same proportions.

- **Locating agent** A substance used to locate, on a chromatogram, the separated parts of a mixture in chromatography.
- **Metals** A class of chemical elements which have a characteristic lustrous appearance and which are good conductors of heat and electricity.
- **Miscible** When two liquids form a homogeneous layer when mixed together, they are said to be miscible.
- **Mixture** A system of two or more substances that can be separated by physical means.
- **Molecule** A group of atoms chemically bonded together.
- **Monatomic molecule** A molecule which consists of only one atom, for example neon and argon.
- **Non-metals** A class of chemical elements that are typically poor conductors of heat and electricity.
- **Oxidation** The process of combining with oxygen.
- **Oxidising agent** A substance which oxidises another substance during a redox reaction.
- **Redox reaction** A reaction which involves the two processes of reduction and oxidation.
- **Reducing agent** A substance which reduces another substance during a redox reaction.
- **Reduction** The process of removing oxygen.
- **Residue** The solid left behind in the filter paper after filtration has taken place.
- **R<sub>f</sub> value** The ratio of the distance travelled by the solute to the distance travelled by the solvent in chromatography.
- **Saturated solution** A solution which contains as much dissolved solute as it can at a particular temperature.
- **Sol** A mixture formed between a solid and a liquid, which then forms a network that can flow.
- **Soluble** If the solute dissolves in the solvent it is said to be soluble.
- **Solute** The substance that dissolves (disappears) in(to) the solvent.
- **Solution** The liquid formed when a substance (solute) disappears (dissolves) into another substance (solvent).
- **Solvent** The liquid that the solute has dissolved in.





# Elements, compounds and experimental techniques

## Additional questions

1 Define the following terms using specific examples to help with your explanation:

- |             |                |
|-------------|----------------|
| a element   | g flocculation |
| b metal     | h gel          |
| c non-metal | i foam         |
| d compound  | j emulsion     |
| e molecule  | k sol.         |
| f mixture   |                |

2 Which of the substances listed below are:

- metallic elements?
- non-metallic elements?
- compounds?
- mixtures?

Silicon, sea water, calcium, argon, water, air, carbon monoxide, iron, sodium chloride, diamond, brass, copper, dilute sulfuric acid, sulfur, oil, nitrogen, ammonia.

3 At room temperature and pressure (rtp), which of the substances listed below is:

- a solid element?
- a liquid element?
- a gaseous mixture?
- a solid mixture?
- a liquid compound?
- a solid compound?

Bromine, carbon dioxide, helium, steel, air, oil, marble, copper, water, sand, tin, bronze, mercury, salt.

4 A student heated a mixture of iron filings and sulfur strongly. He saw a red glow spread through the mixture as the reaction continued. At the end of the experiment a black solid had been formed.

- Explain what the red glow indicates.
- Give the chemical name of the black solid.
- Write a word equation and a balanced chemical equation to represent the reaction which has taken place.
- The black solid is a compound. Explain the difference between the mixture of sulfur and iron and the compound formed by the chemical reaction between them.

5 Name the method which is most suitable for separating the following:

- oxygen from liquid air
- red blood cells from plasma
- petrol and kerosene from crude oil
- coffee grains from coffee solution
- pieces of steel from engine oil
- amino acids from fruit juice solution
- ethanol and water.

6 The table below shows the melting points, boiling points and densities of substances **A** to **D**.

Substance	Melting point/°C	Boiling point/°C	Density/g cm <sup>-3</sup>
A	1110	2606	9.1
B	-266	-252	0.07
C	40	94	1.6
D	-14	60	0.9

- Which substance is a gas at room temperature?
- Which substance is a liquid at room temperature?
- Which substances are solids at room temperature?
- Which substance is most likely to be a metal?
- Which substance will be a liquid at -260 °C?
- What is the melting point of the least dense non-metal?
- Which substances are gases at 72 °C?

7 a How many atoms of the different elements are there in the formulae of the compounds given below?

- nitric acid, HNO<sub>3</sub>
- methane, CH<sub>4</sub>
- copper nitrate, Cu(NO<sub>3</sub>)<sub>2</sub>
- ethanoic acid, CH<sub>3</sub>COOH
- sugar, C<sub>12</sub>H<sub>22</sub>O<sub>11</sub>
- phenol, C<sub>6</sub>H<sub>5</sub>OH
- ammonium sulfate, (NH<sub>4</sub>)<sub>2</sub>SO<sub>4</sub>

b Balance the following equations:

- Zn(s) + O<sub>2</sub>(g) → ZnO(s)
- Fe(s) + Cl<sub>2</sub>(g) → FeCl<sub>3</sub>(s)
- Li(s) + O<sub>2</sub>(g) → Li<sub>2</sub>O(s)
- H<sub>2</sub>(g) + O<sub>2</sub>(g) → H<sub>2</sub>O(g)
- Mg(s) + CO<sub>2</sub>(g) → MgO(s) + C(s)

8 Carbon-fibre-reinforced plastic (CRP) is used in the manufacture of golf clubs and tennis rackets.

a What are composite materials?

b Which two substances are used to manufacture this composite material?

Consider the data in the table.

Material	Strength/ GPa	Stiffness/ GPa	Density/g cm <sup>-3</sup>	Relative cost
Aluminium	0.2	75	2.7	low
Steel	1.1	200	7.8	low
CRP	1.8	195	1.6	high

c Discuss the advantages and disadvantages of using the three materials above in the manufacture of golf clubs.

## 3

# Atomic structure and bonding

## Inside atoms

Proton number and nucleon number  
Ions  
Isotopes  
The mass spectrometer  
Relative atomic mass

## The arrangement of electrons in atoms

### Ionic bonding

Ionic structures  
Properties of ionic compounds  
Formulae of ionic substances  
Oxidation states

### Covalent bonding

Other covalent compounds  
Covalent structures  
Properties of covalent compounds  
Allotropy  
Allotropes of carbon

### Glasses and ceramics

Glasses  
Ceramics

### Metallic bonding

Properties of metals

### Checklist

### Additional questions

We have already seen in Chapter 2 that everything you see around you is made out of tiny particles called atoms (Figure 3.1). When John Dalton developed his atomic theory, about 200 years ago (1807/1808), he stated that the atoms of any one element were identical and that each atom was 'indivisible'. Scientists in those days believed that atoms were solid particles like marbles.

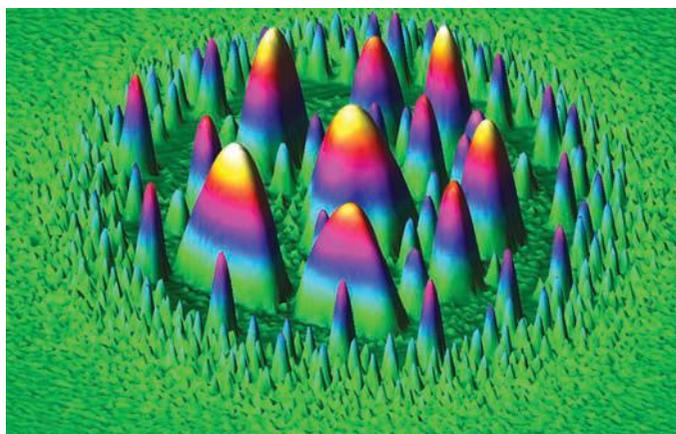


Figure 3.1 A micrograph of atoms.

However, in the last hundred years or so it has been proved by great scientists, such as Niels Bohr, Albert Einstein, Henry Moseley, Joseph Thomson, Ernest Rutherford and James Chadwick, that atoms are in fact made up of even smaller 'sub-atomic' particles. The most important of these are **electrons**, **protons** and **neutrons**, although 70 sub-atomic particles have now been discovered.

## Inside atoms

The three sub-atomic particles are found in distinct and separate regions. The protons and neutrons are found in the centre of the atom, which is called the **nucleus**. The neutrons have no charge and protons are positively charged. The nucleus occupies only a very small volume of the atom but is very dense.

The rest of the atom surrounding the nucleus is where electrons are most likely to be found. The electrons are negatively charged and move around very quickly in **electron shells** or **energy levels**. The electrons are held within the atom by an **electrostatic force of attraction** between themselves and the positive charge of protons in the nucleus (Figure 3.2).

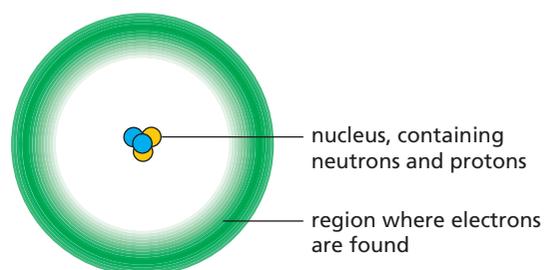


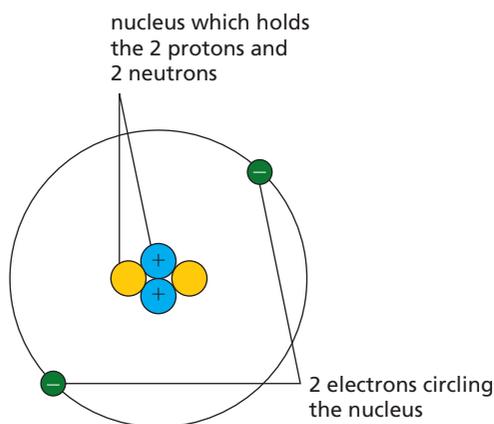
Figure 3.2 Diagram of an atom.

About 1837 electrons are equal in mass to the mass of one proton or one neutron. A summary of each type of particle, its mass and relative charge is shown in Table 3.1. You will notice that the masses of all these particles are measured in **atomic mass units (amu)**. This is because they are so light that their masses cannot be measured usefully in grams.

**Table 3.1** Characteristics of a proton, a neutron and an electron.

Particle	Symbol	Relative mass/amu	Relative charge
Proton	p	1	+1
Neutron	n	1	0
Electron	e	1/1837	-1

Although atoms contain electrically charged particles, the atoms themselves are electrically neutral (they have no overall electric charge). This is because atoms contain equal numbers of electrons and protons. For example, the diagram in Figure 3.3 represents the atom of the non-metallic element helium. The atom of helium possesses two protons, two neutrons and two electrons. The electrical charge of the protons in the nucleus is, therefore, balanced by the opposite charge of the two electrons.

**Figure 3.3** An atom of helium has two protons, two electrons and two neutrons.

## Proton number and nucleon number

The number of protons in the nucleus of an atom is called the **proton number** (or atomic number) and is given the symbol  $Z$ . Hence in the diagram shown in Figure 3.3, the helium atom has a proton number of 2, since it has two protons in its nucleus. Each element has its own proton number and no two different elements have the same proton number. For example, a different element, lithium, has a proton number of 3, since it has three protons in its nucleus.

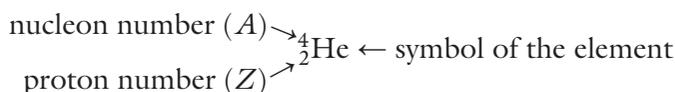
Neutrons and protons have a similar mass. Electrons possess very little mass. So the mass of any atom depends on the number of protons and neutrons in its nucleus. The total number of protons and neutrons found in the nucleus of an atom is called the **nucleon number** (or mass number) and is given the symbol  $A$ .

$$\text{nucleon number} = \text{proton number} + \text{number of neutrons}$$

$$(A) \qquad (Z)$$

Hence, in the example shown in Figure 3.3 the helium atom has a nucleon number of 4, since it has two protons and two neutrons in its nucleus. If we consider the metallic element lithium, it has three protons and four neutrons in its nucleus. It therefore has a nucleon number of 7.

The proton number and nucleon number of an element are usually written in the following shorthand way:



The number of neutrons present can be calculated by rearranging the relationship between the proton number, nucleon number and number of neutrons to give:

$$\begin{array}{l} \text{number of} \\ \text{neutrons} \end{array} = \begin{array}{l} \text{nucleon number} \\ (A) \end{array} - \begin{array}{l} \text{proton number} \\ (Z) \end{array}$$

For example, the number of neutrons in one atom of  ${}_{12}^{24}\text{Mg}$  is:

$$\begin{array}{l} 24 - 12 = 12 \\ (A) \quad (Z) \end{array}$$

and the number of neutrons in one atom of  ${}_{82}^{207}\text{Pb}$  is:

$$\begin{array}{l} 207 - 82 = 125 \\ (A) \quad (Z) \end{array}$$

Table 3.2 shows the number of protons, neutrons and electrons in the atoms of some common elements.

## Ions

An ion is an electrically charged particle. When an atom loses one or more electrons it becomes a positively charged ion. For example, during the chemical reactions of potassium, each atom loses an electron to form a positive ion,  $\text{K}^+$ .

$$\begin{array}{r} {}_{19}\text{K}^+ \\ \begin{array}{r} 19 \text{ protons} = 19+ \\ 18 \text{ electrons} = 18- \\ \hline \text{Overall charge} = 1+ \end{array} \end{array}$$

When an atom gains one or more electrons it becomes a negatively charged ion. For example, during some of the chemical reactions of oxygen it gains two electrons to form a negative ion,  $\text{O}^{2-}$ .

$$\begin{array}{r} {}_8\text{O}^{2-} \\ \begin{array}{r} 8 \text{ protons} = 8+ \\ 10 \text{ electrons} = 10- \\ \hline \text{Overall charge} = 2- \end{array} \end{array}$$

**Table 3.2** Number of protons, neutrons and electrons in some elements.

Element	Symbol	Proton number	Number of electrons	Number of protons	Number of neutrons	Nucleon number
Hydrogen	H	1	1	1	0	1
Helium	He	2	2	2	2	4
Carbon	C	6	6	6	6	12
Nitrogen	N	7	7	7	7	14
Oxygen	O	8	8	8	8	16
Fluorine	F	9	9	9	10	19
Neon	Ne	10	10	10	10	20
Sodium	Na	11	11	11	12	23
Magnesium	Mg	12	12	12	12	24
Sulfur	S	16	16	16	16	32
Potassium	K	19	19	19	20	39
Calcium	Ca	20	20	20	20	40
Iron	Fe	26	26	26	30	56
Zinc	Zn	30	30	30	35	65

Table 3.3 shows some common ions. You will notice from Table 3.3 that:

- some ions contain more than one type of atom, for example  $\text{NO}_3^-$
- an ion may possess more than one unit of charge (either negative or positive), for example  $\text{Al}^{3+}$ ,  $\text{O}^{2-}$  or  $\text{SO}_4^{2-}$ .

**Table 3.3** Some common ions.

Name	Formula
Lithium ion	$\text{Li}^+$
Sodium ion	$\text{Na}^+$
Potassium ion	$\text{K}^+$
Magnesium ion	$\text{Mg}^{2+}$
Calcium ion	$\text{Ca}^{2+}$
Aluminium ion	$\text{Al}^{3+}$
Zinc ion	$\text{Zn}^{2+}$
Ammonium ion	$\text{NH}_4^+$
Fluoride ion	$\text{F}^-$
Chloride ion	$\text{Cl}^-$
Bromide ion	$\text{Br}^-$
Hydroxide ion	$\text{OH}^-$
Oxide ion	$\text{O}^{2-}$
Sulfide ion	$\text{S}^{2-}$
Carbonate ion	$\text{CO}_3^{2-}$
Nitrate ion	$\text{NO}_3^-$
Sulfate ion	$\text{SO}_4^{2-}$

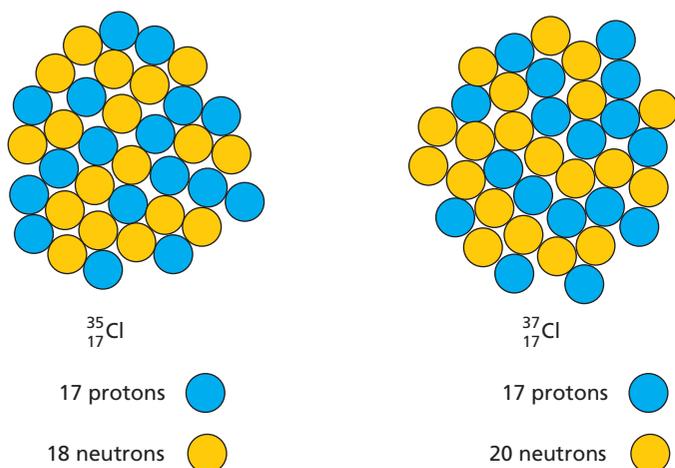
## Isotopes

Not all of the atoms in a sample of chlorine, for example, will be identical. Some atoms of the same element can contain different numbers of neutrons and so have different nucleon numbers. Atoms of the same element which have the same proton number but different neutron numbers are called **isotopes**. The two isotopes of chlorine are shown in Figure 3.4 (p. 36).

Generally, isotopes behave in the same way during chemical reactions because they have the same number of electrons on their outer shell (see p. 38).

The only effect of the extra neutron is to alter the mass of the atom and properties which depend on it, such as density. Some other examples of atoms with isotopes are shown in Table 3.4.

There are two types of isotopes: those which are stable and those which are unstable. The isotopes which are unstable, as a result of the extra neutrons in their nuclei, are **radioactive** and are called **radioisotopes**. For example, uranium-235, which is used as a source of power in nuclear reactors (p. 93), and cobalt-60, which is used in radiotherapy treatment (Figure 3.5), are both radioisotopes.



**Figure 3.4** The two isotopes of chlorine.

**Table 3.4** Some atoms and their isotopes.

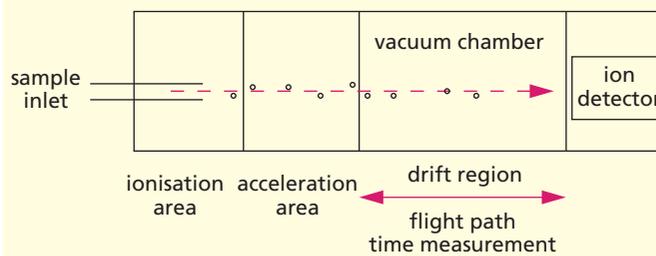
Element	Symbol	Particles present
Hydrogen	$^1_1\text{H}$	1 e, 1 p, 0 n
(Deuterium)	$^2_1\text{H}$	1 e, 1 p, 1 n
(Tritium)	$^3_1\text{H}$	1 e, 1 p, 2 n
Carbon	$^{12}_6\text{C}$	6 e, 6 p, 6 n
	$^{13}_6\text{C}$	6 e, 6 p, 7 n
	$^{14}_6\text{C}$	6 e, 6 p, 8 n
Oxygen	$^{16}_8\text{O}$	8 e, 8 p, 8 n
	$^{17}_8\text{O}$	8 e, 8 p, 9 n
	$^{18}_8\text{O}$	8 e, 8 p, 10 n
Strontium	$^{86}_{38}\text{Sr}$	38 e, 38 p, 48 n
	$^{88}_{38}\text{Sr}$	38 e, 38 p, 50 n
	$^{90}_{38}\text{Sr}$	38 e, 38 p, 52 n
Uranium	$^{235}_{92}\text{U}$	92 e, 92 p, 143 n
	$^{238}_{92}\text{U}$	92 e, 92 p, 146 n



**Figure 3.5** Cobalt-60 is used in radiotherapy treatment.

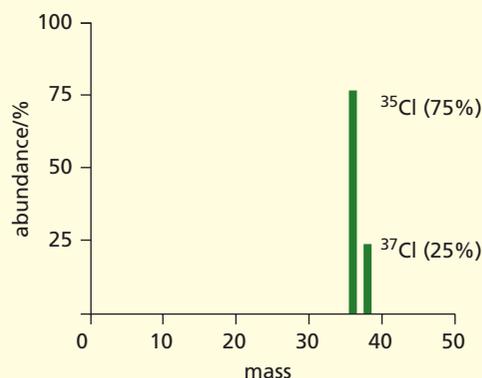
## The mass spectrometer

How do we know isotopes exist? They were first discovered by scientists using apparatus called a mass spectrometer (Figure 3.6). The first mass spectrometer was built by the British scientist Francis Aston in 1919 and enabled scientists to compare the relative masses of atoms accurately for the first time.



**Figure 3.6** A diagram of a mass spectrometer.

A vacuum exists inside the mass spectrometer. A sample of the vapour of the element is injected into the ionisation chamber where it is bombarded by electrons. The collisions which take place between these electrons and the injected atoms cause an electron to be lost from the atom, which becomes a positive ion with a +1 charge. These positive ions are then accelerated towards a negatively charged plate, in the acceleration area. The spectrometer is set up to ensure that when the ions leave the acceleration area they all have the same kinetic energy, regardless of the mass of the ions. This means that the lighter ions travel faster than the heavier ones, and effectively separates the ions according to their mass. Having left the acceleration area, the time for the ions to reach the detector is recorded. The detector counts the number of each of the ions which fall upon it and so a measure of the percentage abundance of each isotope is obtained. A typical mass spectrum for chlorine is shown in Figure 3.7.



**Figure 3.7** The mass spectrum for chlorine.

## Relative atomic mass

The average mass of a large number of atoms of an element is called its **relative atomic mass** (symbol  $A_r$ ). This quantity takes into account the percentage abundance of all the isotopes of an element which exist.

In 1961 the International Union of Pure and Applied Chemistry (IUPAC) recommended that the standard used for the  $A_r$  scale was carbon-12. An atom of carbon-12 was taken to have a mass of 12 amu. The  $A_r$  of an element is the average mass of the naturally occurring atoms of an element on a scale where  $^{12}\text{C}$  has a mass of exactly 12 units:

$$A_r = \frac{\text{average mass of isotopes of the element}}{\frac{1}{12} \times \text{mass of 1 atom of carbon-12}}$$

Note:  $\frac{1}{12}$  of the mass of one carbon-12 atom = 1 amu.  
For example, chlorine has two isotopes:

	$^{35}_{17}\text{Cl}$	$^{37}_{17}\text{Cl}$
% abundance	75	25

Hence the 'average mass' or  $A_r$  of a chlorine atom is:

$$\frac{(75 \times 35) + (25 \times 37)}{100} = 35.5$$

$$A_r = \frac{35.5}{1} = 35.5 \text{ amu}$$

### Questions

- Calculate the number of neutrons in the following atoms:  
 a  $^{27}_{13}\text{Al}$     b  $^{31}_{15}\text{P}$     c  $^{262}_{107}\text{Uns}$     d  $^{190}_{76}\text{Os}$
- Given that the percentage abundance of  $^{20}_{10}\text{Ne}$  is 90% and that of  $^{22}_{10}\text{Ne}$  is 10%, calculate the  $A_r$  of neon.

## ● The arrangement of electrons in atoms

The nucleus of an atom contains the heavier sub-atomic particles – the protons and the neutrons. The electrons, the lightest of the sub-atomic particles, move around the nucleus at great distances from the nucleus relative to their size. They move very fast in electron energy levels very much as the planets orbit the Sun.

It is not possible to give the exact position of an electron in an energy level. However, we can state that electrons can only occupy certain, definite energy levels and that they cannot exist between them. Each of the electron energy levels can hold only a certain number of electrons.

- First energy level holds up to two electrons.
- Second energy level holds up to eight electrons.
- Third energy level holds up to 18 electrons.

There are further energy levels which contain increasing numbers of electrons.

The third energy level can be occupied by a maximum of 18 electrons. However, when eight electrons have occupied this level a certain stability is given to the atom and the next two electrons go into the fourth energy level, and then the remaining ten electrons complete the third energy level.

The electrons fill the energy levels starting from the energy level nearest to the nucleus, which has the lowest energy. When this is full (with two electrons) the next electron goes into the second energy level. When this energy level is full with eight electrons, then the electrons begin to fill the third and fourth energy levels as stated above.

For example, a  $^{16}_8\text{O}$  atom has a proton number of 8 and therefore has eight electrons. Two of the eight electrons enter the first energy level, leaving six to occupy the second energy level, as shown in Figure 3.8. The electron configuration for oxygen can be written in a shorthand way as 2,6.

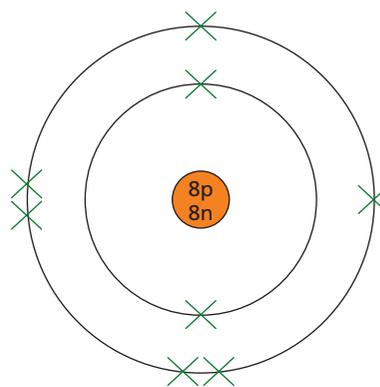


Figure 3.8 Arrangement of electrons in an oxygen atom.

There are 118 elements, and Table 3.5 shows the way in which the electrons are arranged in the first 20 of these elements. The way in which the electrons are distributed is called the **electronic structure** (or electron configuration). Figure 3.9 shows the electronic structure of a selection of atoms.

**Table 3.5** Electronic structure of the first 20 elements.

Element	Symbol	Proton number	Number of electrons	Electron structure
Hydrogen	H	1	1	1
Helium	He	2	2	2
Lithium	Li	3	3	2,1
Beryllium	Be	4	4	2,2
Boron	B	5	5	2,3
Carbon	C	6	6	2,4
Nitrogen	N	7	7	2,5
Oxygen	O	8	8	2,6
Fluorine	F	9	9	2,7
Neon	Ne	10	10	2,8
Sodium	Na	11	11	2,8,1
Magnesium	Mg	12	12	2,8,2
Aluminium	Al	13	13	2,8,3
Silicon	Si	14	14	2,8,4
Phosphorus	P	15	15	2,8,5
Sulfur	S	16	16	2,8,6
Chlorine	Cl	17	17	2,8,7
Argon	Ar	18	18	2,8,8
Potassium	K	19	19	2,8,8,1
Calcium	Ca	20	20	2,8,8,2

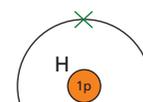
You will notice from Table 3.5 that the elements helium, neon and argon have completely full outer shells. In Chapter 9 you will see that these elements are known as the noble or inert gases and that they are generally very stable and unreactive (p. 143). This is linked to the full outer shells that they possess. It would seem that when elements react to form compounds they do so to achieve full electron energy levels. This idea forms the basis of the electronic theory of chemical bonding.

### Questions

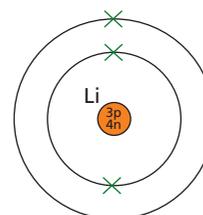
- How many electrons may be accommodated in the first three energy levels?
- What is the same about the electron structures of:
  - lithium, sodium and potassium?
  - beryllium, magnesium and calcium?

## ● Ionic bonding

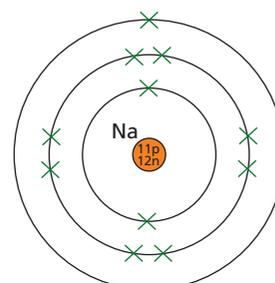
Ionic bonds are usually found in compounds that contain metals combined with non-metals. When this type of bond is formed, electrons are transferred from the metal atoms to the non-metal atoms during the



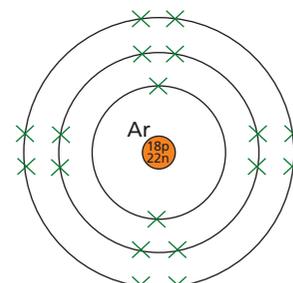
hydrogen



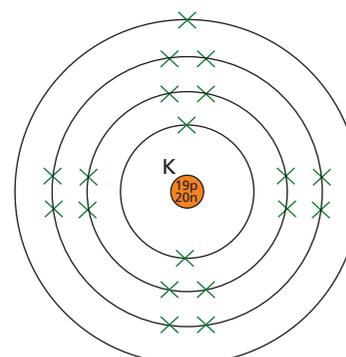
lithium



sodium



argon



potassium

**Figure 3.9** Electron arrangements of hydrogen, lithium, sodium, argon and potassium.

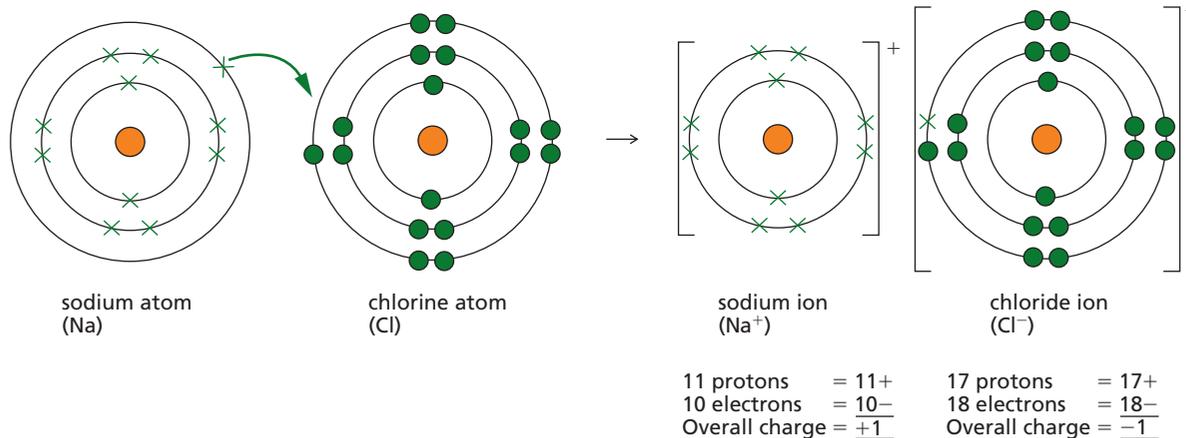
chemical reaction. In doing this, the atoms become more stable by getting full outer energy levels. For example, consider what happens when sodium and chlorine react together and combine to make sodium chloride (Figure 3.10).

sodium + chlorine  $\rightarrow$  sodium chloride



**Figure 3.10** The properties of salt are very different from those of the sodium and chlorine it was made from. To get your salt you would not eat sodium or inhale chlorine!

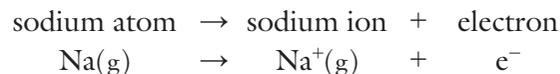
Sodium has just one electron in its outer energy level ( $_{11}\text{Na}$  2,8,1). Chlorine has seven electrons in its outer energy level ( $_{17}\text{Cl}$  2,8,7). When these two elements react,



**Figure 3.11** Ionic bonding in sodium chloride.

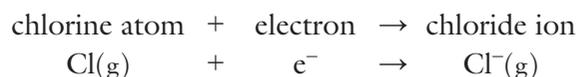
the outer electron of each sodium atom is transferred to the outer energy level of a chlorine atom (Figure 3.11).

In this way both the atoms obtain full outer energy levels and become 'like' the nearest noble gas. The sodium atom has become a sodium ion and the process is known as **ionisation**. This sodium ion has an electron configuration like neon.



To lose electrons in this way is called **oxidation**.

The chlorine atom has become a chloride ion with an electron configuration like argon.



To gain electrons in this way is called **reduction**.

In the chemical process producing sodium chloride both oxidation and reduction have taken place and so this is known as a **redox reaction**. A further discussion of oxidation and reduction in terms of electron transfer takes place in Chapter 5 (p. 73).

Only the outer electrons are important in bonding, so we can simplify the diagrams by missing out the inner energy levels (Figure 3.12, p. 40).

The charges on the sodium and chloride ions are equal but opposite. They balance each other and the resulting formula for sodium chloride is NaCl. These oppositely charged ions attract each other and are pulled, or **bonded**, to one another by strong electrostatic forces. This type of bonding is called **ionic bonding**. The alternative name, **electrovalent bonding**, is derived from the fact that there are electrical charges on the atoms involved in the bonding.

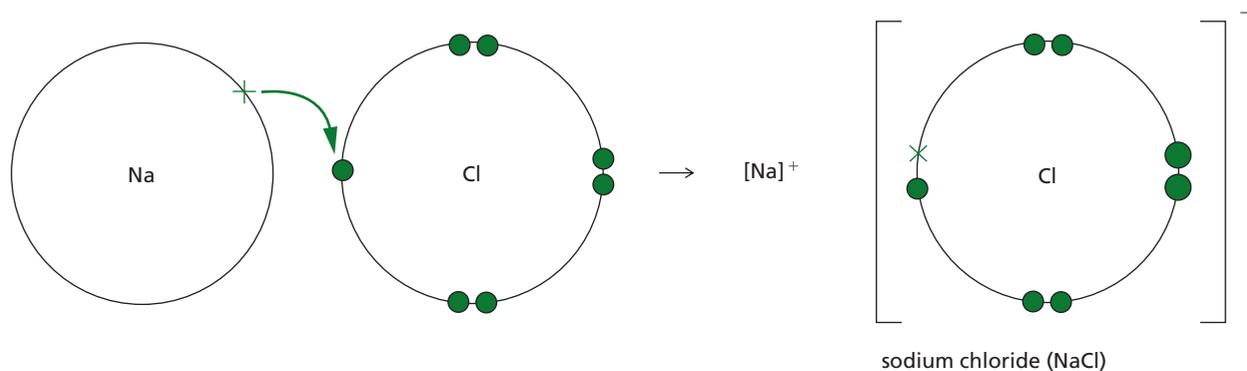


Figure 3.12 Simplified diagram of ionic bonding in sodium chloride.

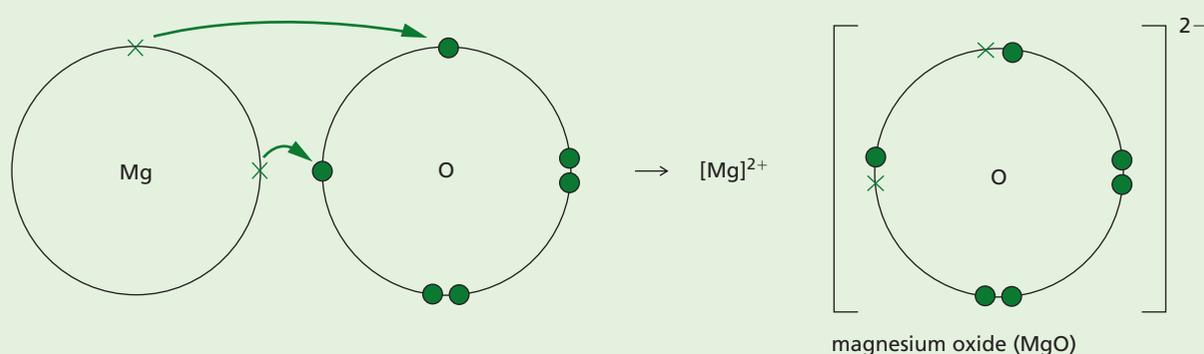


Figure 3.13 Simplified diagram of ionic bonding in magnesium oxide.

Figure 3.13 shows the electron transfers that take place between a magnesium atom and an oxygen atom during the formation of magnesium oxide.

Magnesium obtains a full outer energy level by losing two electrons. These are transferred to the oxygen atom. In magnesium oxide, the  $\text{Mg}^{2+}$  and  $\text{O}^{2-}$  are oppositely charged and are attracted to one another. The formula for magnesium oxide is  $\text{MgO}$ .

Figure 3.14 shows the electron transfers that take place during the formation of calcium chloride. When these two elements react, the calcium atom gives each of the two chlorine atoms one electron.

In this case, a compound is formed containing two chloride ions ( $\text{Cl}^-$ ) for each calcium ion ( $\text{Ca}^{2+}$ ). The chemical formula is  $\text{CaCl}_2$ .

### Question

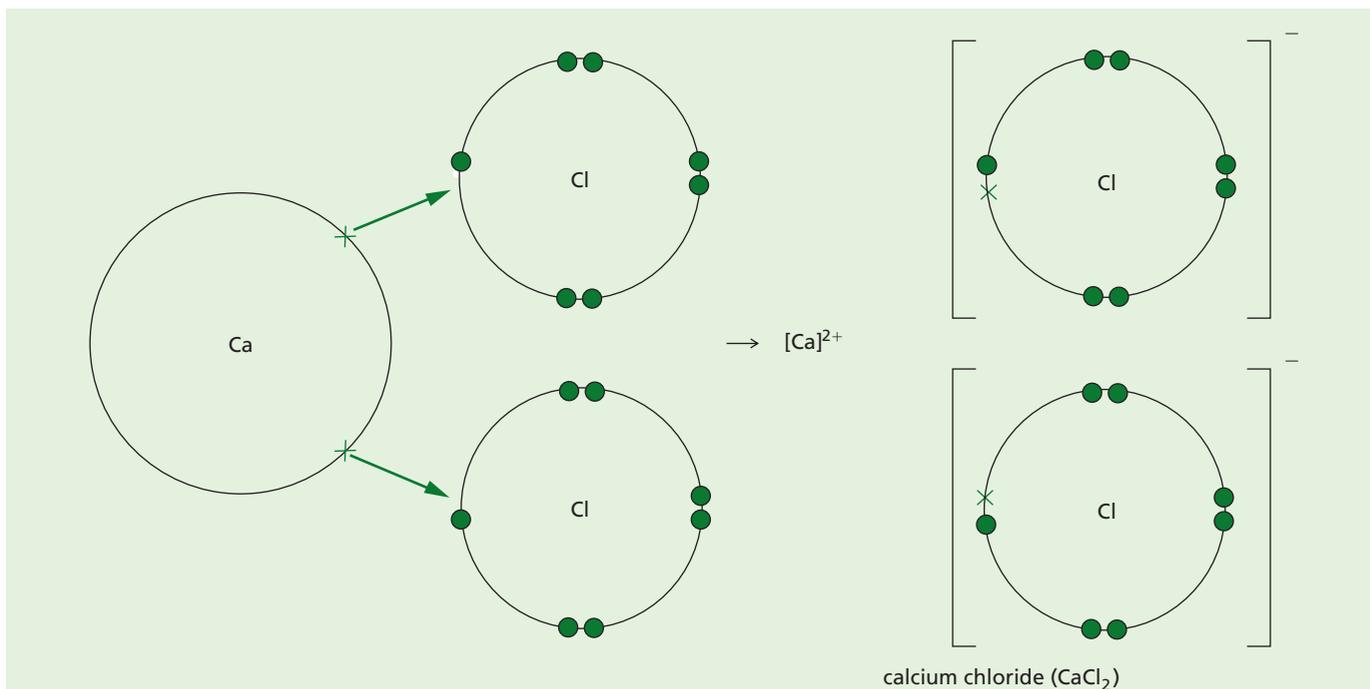
- 1 Draw diagrams to represent the bonding in each of the following ionic compounds:
  - a potassium fluoride (KF)
  - b lithium chloride (LiCl)
  - c magnesium fluoride ( $\text{MgF}_2$ )
  - d calcium oxide (CaO).

Scientists, using **X-ray diffraction** (Figure 3.15a), have obtained photographs that indicate the way in which the ions are arranged (Figure 3.15b). The electron density map of sodium chloride is shown in Figure 3.15c.

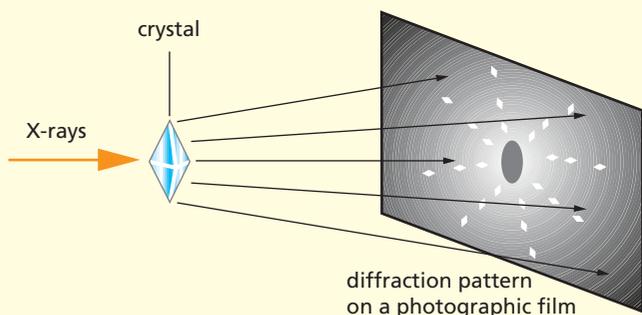
Figure 3.15d shows the structure of sodium chloride as determined by the X-ray diffraction technique. The study of crystals using X-ray diffraction was pioneered by Sir William Bragg and his son Sir Lawrence Bragg in 1912. X-rays are a form of electromagnetic radiation. They have a much

shorter wavelength than light therefore it is possible to use them to investigate extremely small structures.

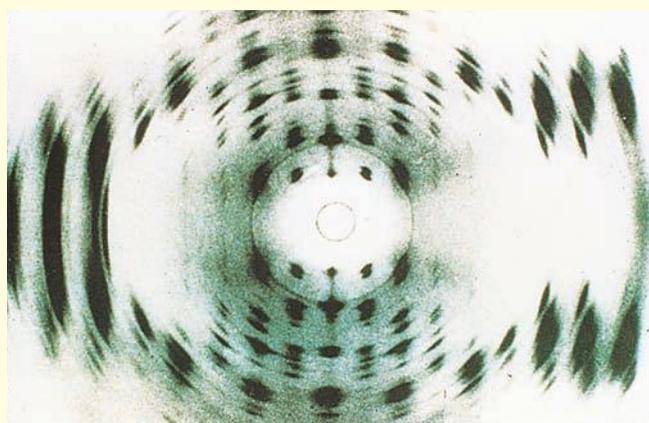
When X-rays are passed through a crystal of sodium chloride, for example, you get a pattern of spots called a diffraction pattern (Figure 3.15b). This pattern can be recorded on photographic film and used to work out how the ions or atoms are arranged in the crystal. Crystals give particular diffraction patterns depending on their structure, and this makes X-ray diffraction a particularly powerful technique in the investigation of crystal structures.



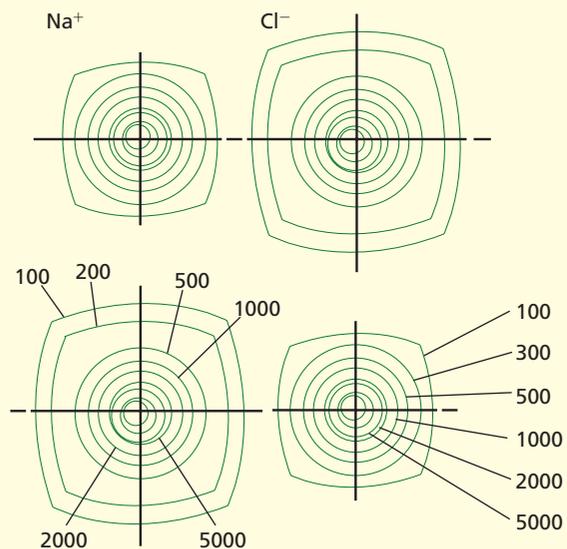
**Figure 3.14** The transfer of electrons that takes place during the formation of calcium chloride.



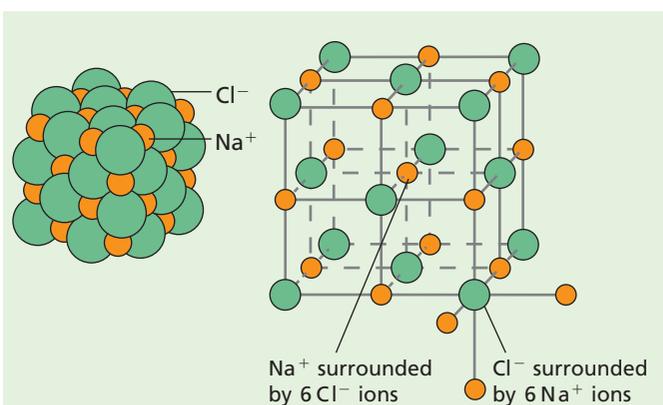
**a** X-ray diffraction technique.



**b** X-ray diffraction photograph of sodium chloride.



**c** Electron density map of sodium chloride.



**d** The structure of sodium chloride.

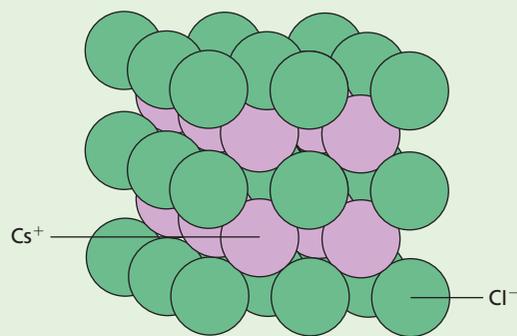
**Figure 3.15**

## Ionic structures

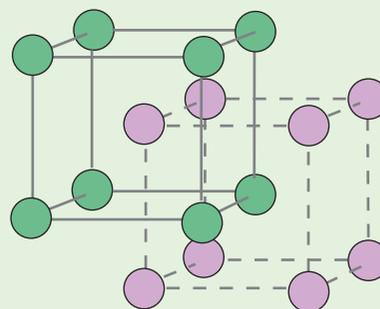
Ionic structures are solids at room temperature and have high melting and boiling points. The ions are packed together in a regular arrangement called a **lattice**. Within the lattice, oppositely charged ions attract one another strongly.

Figure 3.15d shows only a tiny part of a small crystal of sodium chloride. Many millions of sodium ions and chloride ions would be arranged in this way in a crystal of sodium chloride to make up the giant ionic structure. Each sodium ion in the lattice is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions.

Not all ionic substances form the same structures. Caesium chloride ( $\text{CsCl}$ ), for example, forms a different structure due to the larger size of the caesium ion compared with that of the sodium ion. This gives rise to the structure shown in Figure 3.16, which is called a body-centred cubic structure. Each caesium ion is surrounded by eight chloride ions and, in turn, each chloride ion is surrounded by eight caesium ions.



a The arrangement of ions.



$\text{Cs}^+$  surrounded by 8  $\text{Cl}^-$  ions  
In the same way, each  $\text{Cl}^-$  is surrounded by 8  $\text{Cs}^+$  ions

b The unit cell.

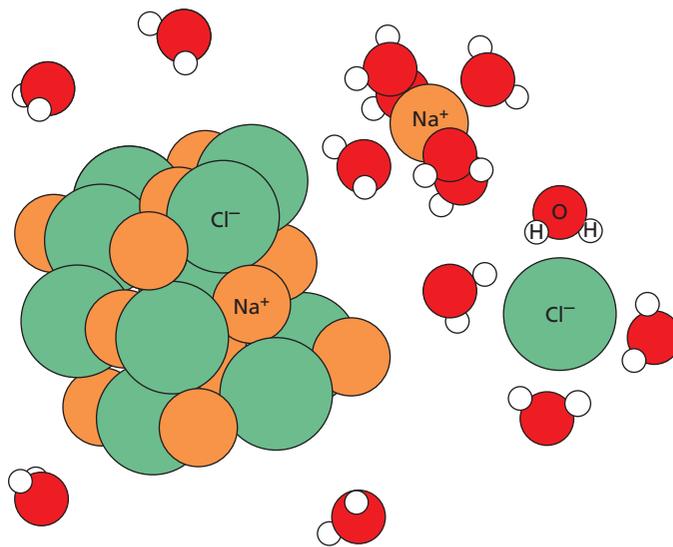
**Figure 3.16** The structure of caesium chloride.

## Properties of ionic compounds

Ionic compounds have the following properties.

- They are usually solids at room temperature, with high melting points. This is due to the strong electrostatic forces holding the crystal lattice together. A lot of energy is therefore needed to separate the ions and melt the substance.
- They are usually hard substances.
- They usually cannot conduct electricity when solid, because the ions are not free to move.
- They mainly dissolve in water. This is because water molecules are able to bond with both the positive and the negative ions, which breaks up the lattice and keeps the ions apart. Figure 3.17 shows the interaction between water molecules (the solvent) and sodium and chloride ions from sodium chloride (the solute). For a further discussion of the solubility of substances see Chapter 8 (p. 129).
- They usually conduct electricity when in the molten state or in aqueous solution. The forces of attraction between the ions are weakened and the

ions are free to move to the appropriate electrode. This allows an electric current to be passed through the molten compound (see Chapter 5, p. 73).



**Figure 3.17** Salt (an ionic compound) dissolving in water.

## Formulae of ionic substances

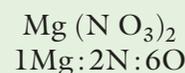
On p. 39 we saw that ionic compounds contain positive and negative ions, whose charges balance. For example, sodium chloride contains one  $\text{Na}^+$  ion for every  $\text{Cl}^-$  ion, giving rise to the formula  $\text{NaCl}$ . This method can be used to write down formulae which show the ratio of the number of ions present in any ionic compound.

The formula of magnesium chloride is  $\text{MgCl}_2$ . This formula is arrived at by each  $\text{Mg}^{2+}$  ion combining with two  $\text{Cl}^-$  ions, and once again the charges balance. The size of the charge on an ion is a measure of its **valency** or **combining power**.  $\text{Na}^+$  has a valency of 1, but  $\text{Mg}^{2+}$  has a valency of 2.  $\text{Na}^+$  can bond (combine) with only one  $\text{Cl}^-$  ion, whereas  $\text{Mg}^{2+}$  can bond with two  $\text{Cl}^-$  ions.

Some elements, such as copper and iron, possess two ions with different valencies. Copper can form the  $\text{Cu}^+$  ion and the  $\text{Cu}^{2+}$  ion, with valencies 1 and 2 respectively. Therefore it can form two different compounds with chlorine,  $\text{CuCl}$  and  $\text{CuCl}_2$ . We can also distinguish the difference by using Roman numerals in their names:  $\text{CuCl}$  is copper(I) chloride and  $\text{CuCl}_2$  is copper(II) chloride. Similarly, iron forms the  $\text{Fe}^{2+}$  and  $\text{Fe}^{3+}$  ions and so can also form two different compounds with, for example, chlorine:  $\text{FeCl}_2$  (iron(II) chloride) and  $\text{FeCl}_3$  (iron(III) chloride).

Table 3.6 shows the valencies of a series of ions you will normally meet in your study of chemistry.

You will notice that Table 3.6 includes groups of atoms which have net charges. For example, the nitrate ion is a single unit composed of one nitrogen atom and three oxygen atoms and has one single negative charge. The formula, therefore, of magnesium nitrate would be  $\text{Mg}(\text{NO}_3)_2$ . You will notice that the  $\text{NO}_3$  has been placed in brackets with a 2 outside the bracket. This indicates that there are two nitrate ions present for every magnesium ion. The ratio of the atoms present is therefore:



The charge on the element ion is often referred to as its **oxidation state**.

## Oxidation states

Each atom in an element or compound is assigned an oxidation state to show how much it is reduced or oxidised. The following points should be remembered when using oxidation states.

- The oxidation state of the free element is always 0, for example in metals such as zinc and copper.
- In simple ions, the oxidation state is the same as the charge on the ion. So iodine has an oxidation state of 0 in  $\text{I}_2$  but an oxidation state of  $-1$  in  $\text{I}^-$ .

**Table 3.6** Valencies (and oxidation states) of some elements (ions) and groups of atoms.

	Valency (oxidation state)					
	1		2		3	
Metals	Lithium	( $\text{Li}^+$ )	Magnesium	( $\text{Mg}^{2+}$ )	Aluminium	( $\text{Al}^{3+}$ )
	Sodium	( $\text{Na}^+$ )	Calcium	( $\text{Ca}^{2+}$ )	Iron(III)	( $\text{Fe}^{3+}$ )
	Potassium	( $\text{K}^+$ )	Copper(II)	( $\text{Cu}^{2+}$ )		
	Silver	( $\text{Ag}^+$ )	Zinc	( $\text{Zn}^{2+}$ )		
	Copper(I)	( $\text{Cu}^+$ )	Iron(II)	( $\text{Fe}^{2+}$ )		
			Lead	( $\text{Pb}^{2+}$ )		
		Barium	( $\text{Ba}^{2+}$ )			
Non-metals	Fluoride	( $\text{F}^-$ )	Oxide	( $\text{O}^{2-}$ )		
	Chloride	( $\text{Cl}^-$ )	Sulfide	( $\text{S}^{2-}$ )		
	Bromide	( $\text{Br}^-$ )				
	Hydrogen	( $\text{H}^+$ )				
Groups of atoms	Hydroxide	( $\text{OH}^-$ )	Carbonate	( $\text{CO}_3^{2-}$ )	Phosphate	( $\text{PO}_4^{3-}$ )
	Nitrate	( $\text{NO}_3^-$ )	Sulfate	( $\text{SO}_4^{2-}$ )		
	Ammonium	( $\text{NH}_4^+$ )	Dichromate(VI)	( $\text{Cr}_2\text{O}_7^{2-}$ )		
	Hydrogencarbonate	( $\text{HCO}_3^-$ )				
	Manganate(VII)	( $\text{MnO}_4^-$ )				

- Compounds have no charge overall. Hence the oxidation states of all the individual elements in a compound must add up to 0. The oxidation states of elements in compounds can vary, as seen in Table 3.6. It is possible to recognise which of the different oxidation states a metal element is in by the colour of its compounds (Figure 3.18).



Figure 3.18 Iron(II) sulfate is pale green, whilst iron(III) sulfate is yellow.

- An increase in the oxidation state, for example from +2 to +3 as in the case of  $\text{Fe}^{2+}$  to  $\text{Fe}^{3+}$ , is oxidation.
- However, a reduction in the oxidation state, for example from +6 to +3 as in the case of  $\text{Cr}^{6+}$  (in  $\text{CrO}_4^{2-}$ ) to  $\text{Cr}^{3+}$ , is reduction.

During a redox reaction, the substance that brings about oxidation is called an **oxidising agent** and is itself reduced during the process. A substance that brings about reduction is a **reducing agent** and is itself oxidised during the process (see pp. 14, 39).

For example, if a dilute solution of acidified potassium manganate(VII) (pale purple) is added to a solution of iron(II) sulfate, a colour change takes place as the reaction occurs (Figure 3.19). The iron(II) sulfate (pale green) changes to pale yellow, showing the presence of iron(III) ions.



Figure 3.19 Manganate(VII) ions (oxidising agent) and iron(II) ions (reducing agent) are involved in a redox reaction when mixed.

In this reaction the iron(II) ions have been oxidised to iron(III) ions (increase in oxidation state) and the manganate(VII) ions have been reduced to manganese(II) ions (decrease in oxidation state) which are very pale pink. Hence the manganate(VII) ions are the oxidising agent and the iron(II) ions are the reducing agent.

It is possible to recognise redox processes by looking at the oxidation states on the two sides of the chemical equation for a reaction. For example, in the reaction between magnesium and dilute sulfuric acid, the magnesium dissolves and hydrogen gas is produced. Both magnesium metal and hydrogen gas are free elements and so have an oxidation state of 0. In sulfuric acid, hydrogen has an oxidation state of +1 since the overall charge on the sulfate ion is  $-2$ . Similarly, the oxidation state of magnesium in magnesium sulfate is +2.

magnesium + sulfuric acid  $\rightarrow$  magnesium sulfate + hydrogen



Oxidation states: 0 +1 +2 0

The sulfate ions are unchanged by the reaction and so can be ignored.

As you can see, the oxidation state of magnesium has increased from 0 to +2. Therefore the magnesium has been oxidised by the sulfuric acid and so sulfuric acid is the oxidising agent. The oxidation state of hydrogen has decreased from +1 in the sulfuric acid to 0 in the free element. Therefore the hydrogen has been reduced by the magnesium and so magnesium is the reducing agent.

### Question

1 Identify the oxidising and reducing agents in the following reactions.

- $\text{Zn(s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{ZnSO}_4(\text{aq}) + \text{H}_2(\text{g})$
- $\text{Cu}^{2+}(\text{aq}) + \text{Zn(s)} \rightarrow \text{Cu(s)} + \text{Zn}^{2+}(\text{aq})$
- $\text{Mg(s)} + \text{Cu}(\text{NO}_3)_2(\text{aq}) \rightarrow \text{Mg}(\text{NO}_3)_2(\text{aq}) + \text{Cu(s)}$
- $\text{Fe(s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{FeSO}_4(\text{aq}) + \text{H}_2(\text{g})$

### The 'cross-over' method

A less scientific but simpler method to work out the formula of compounds is called the 'cross-over' method. In this method it is only necessary to 'swap' the valencies of the elements or groups of atoms

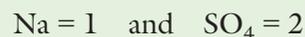
(or radicals) concerned. This is easily done by 'crossing over' their valencies, by placing the valency of the first element after the symbol of the second and placing the valency of the second element or radical after the symbol of the first.

For example, in aluminium sulfide the valencies of the elements are:



Hence the chemical formula is  $\text{Al}_2\text{S}_3$ .

Similarly, in sodium sulfate the valencies are:



Hence the chemical formula is  $\text{Na}_2\text{SO}_4$ .

### Questions

- Using the information in Table 3.6, write the formulae for:
 

a copper(I) oxide	d lead bromide
b zinc phosphate	e potassium manganate(VII)
c iron(III) chloride	f sodium dichromate(VI)
- Using the formulae in your answer to question 1, write down the ratio of atoms present for each of the compounds.

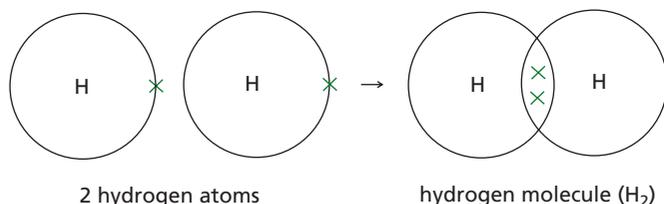
## ● Covalent bonding

Another way in which atoms can gain the stability of the noble gas electron configuration is by sharing the electrons in their outer energy levels. This occurs between non-metal atoms, and the bond formed is called a **covalent bond**. The simplest example of this type of bonding can be seen by considering the hydrogen molecule,  $\text{H}_2$ .

Each hydrogen atom in the molecule has one electron. In order to obtain a full outer energy level and gain the electron configuration of the noble gas helium, each of the hydrogen atoms must have two electrons. To do this, the two hydrogen atoms allow their outer energy levels to overlap (Figure 3.20a). A molecule of hydrogen is formed, with two hydrogen atoms sharing a pair of electrons (Figure 3.20a, b, c). This shared pair of electrons is known as a single covalent bond and is represented by a single line as in hydrogen:



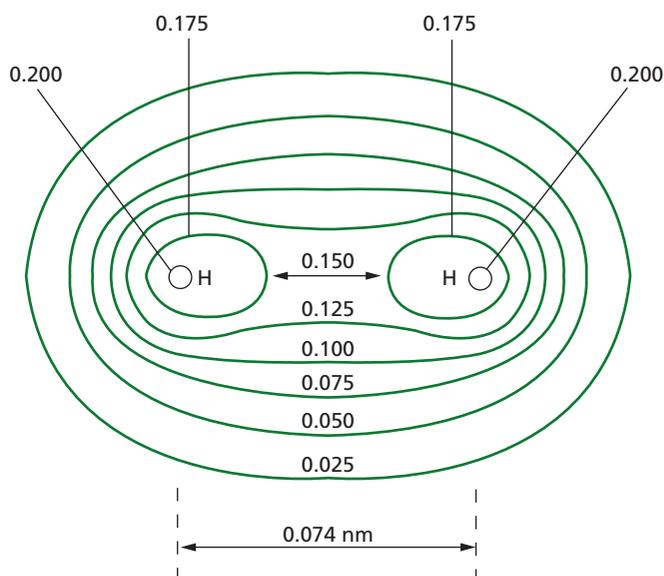
A similar example exists in the diatomic halogen molecule chlorine,  $\text{Cl}_2$  (Figure 3.21, p. 46).



**Figure 3.20a** The electron sharing to form the single covalent bond in  $\text{H}_2$  molecules.

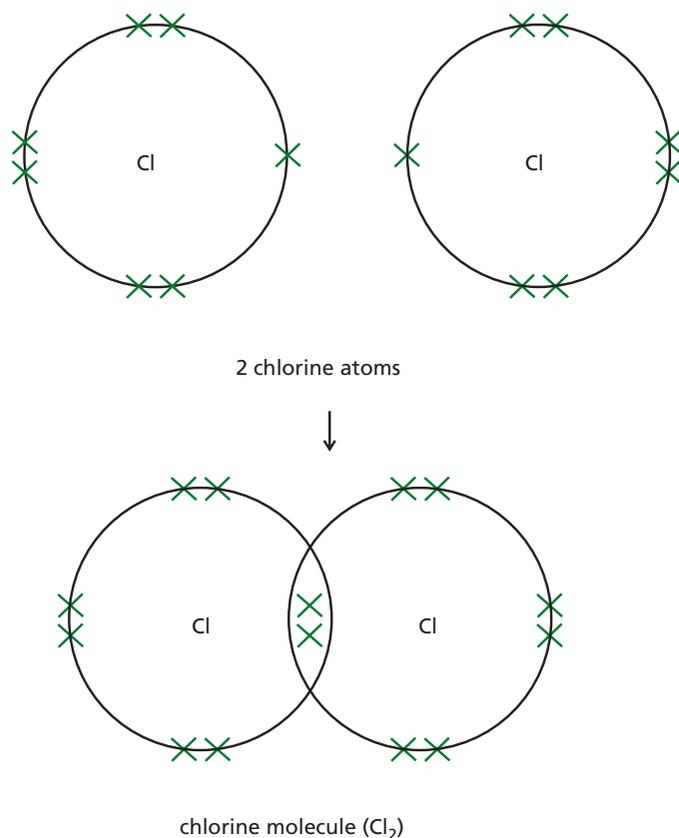


**Figure 3.20b** Model of hydrogen molecule.



electron densities in electrons per cubic atomic unit of length

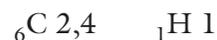
**Figure 3.20c** Electron density map of a hydrogen molecule.



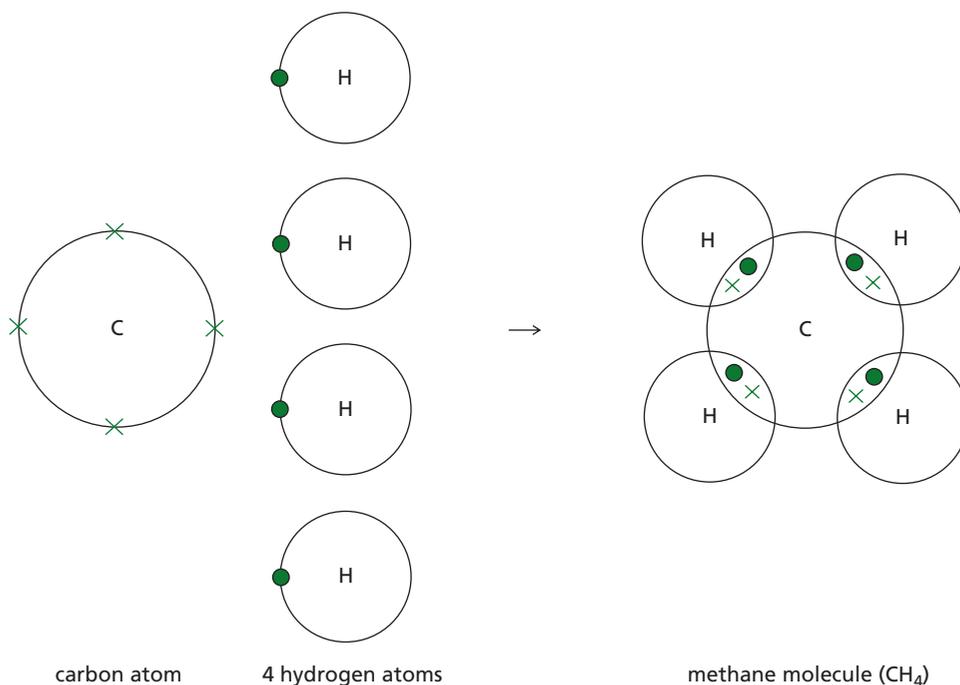
**Figure 3.21** The electron sharing to form the single covalent bond in  $\text{Cl}_2$  molecules (sharing outer electron shells only).

## Other covalent compounds

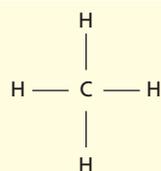
**Methane** (natural gas) is a gas whose molecules contain atoms of carbon and hydrogen. The electron structures are:



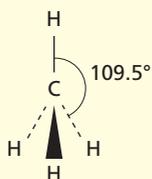
The carbon atom needs four more electrons to attain the electron configuration of the noble gas neon. Each hydrogen atom needs only one electron to form the electron configuration of helium. Figure 3.22 shows how the atoms gain these electron configurations by the sharing of electrons. You will note that only the outer electron energy levels are shown. Figure 3.23 shows the shape of the methane molecule.



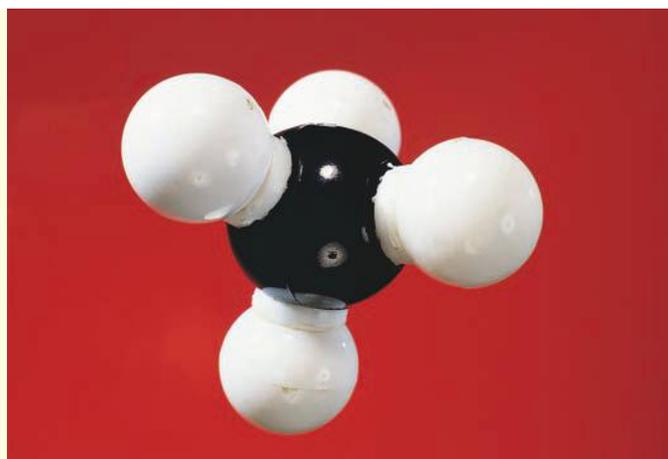
**Figure 3.22** Formation of methane.



the methane molecule is tetrahedral



a Methane molecule.



b Model of the methane molecule. The tetrahedral shape is caused by the repulsion of the C–H bonding pairs of electrons.

Figure 3.23

**Ammonia** is a gas containing the elements nitrogen and hydrogen. It is used in large amounts to make fertilisers. The electron configurations of the two elements are:



The nitrogen atom needs three more electrons to obtain the noble gas structure of neon. Each hydrogen requires only one electron to form the noble gas structure of helium. The nitrogen and hydrogen atoms share electrons, forming three single covalent bonds (Figure 3.24). Unlike methane the shape of an ammonia molecule is pyramidal (Figure 3.25, p. 48).

**Water** is a liquid containing the elements hydrogen and oxygen. The electronic structures of the two elements are:



The oxygen atom needs two electrons to gain the electron configuration of neon. Each hydrogen requires one more electron to gain the electron configuration of helium. Again, the oxygen and hydrogen atoms share electrons, forming a water molecule with two single covalent bonds as shown in Figure 3.26. A water molecule is V-shaped (Figure 3.27).

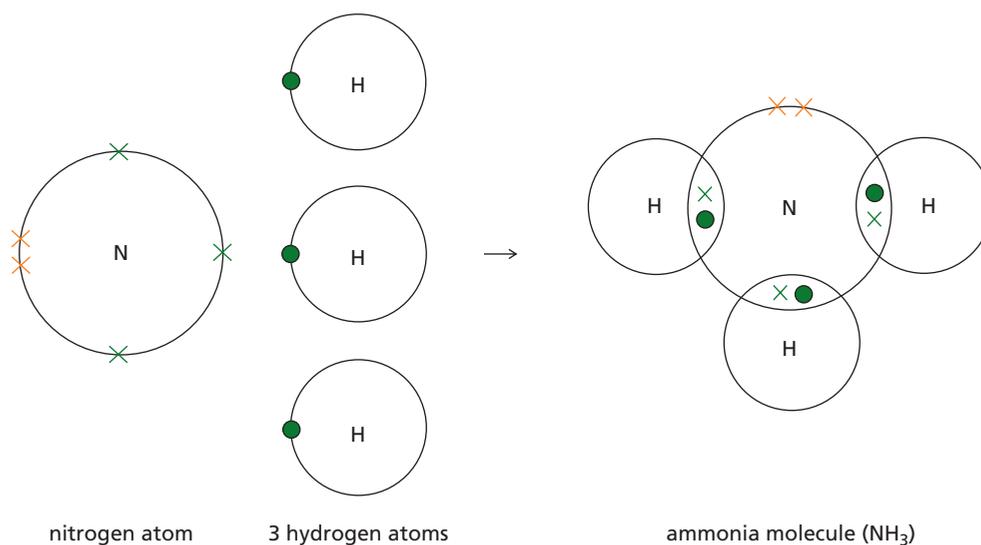
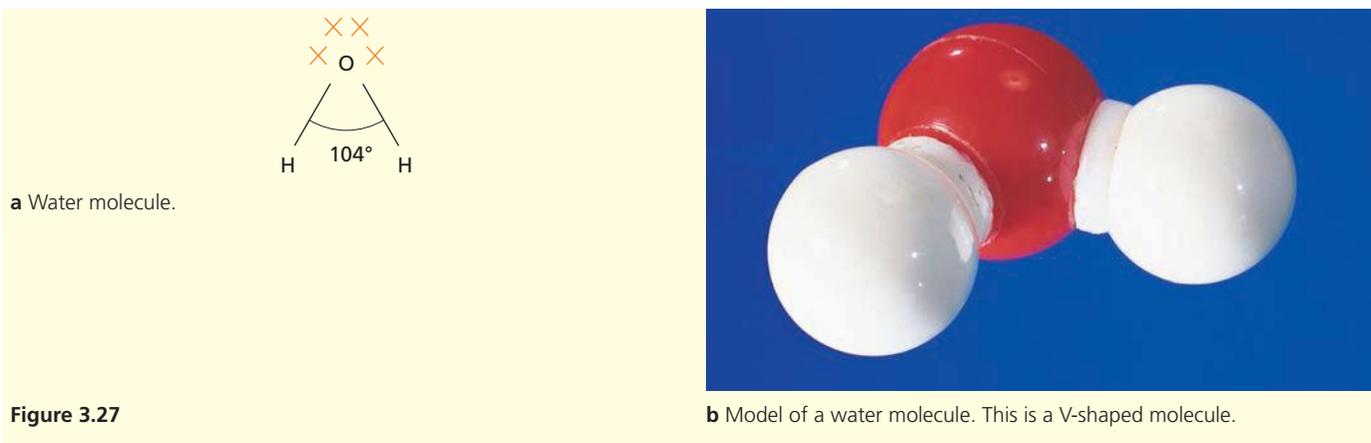
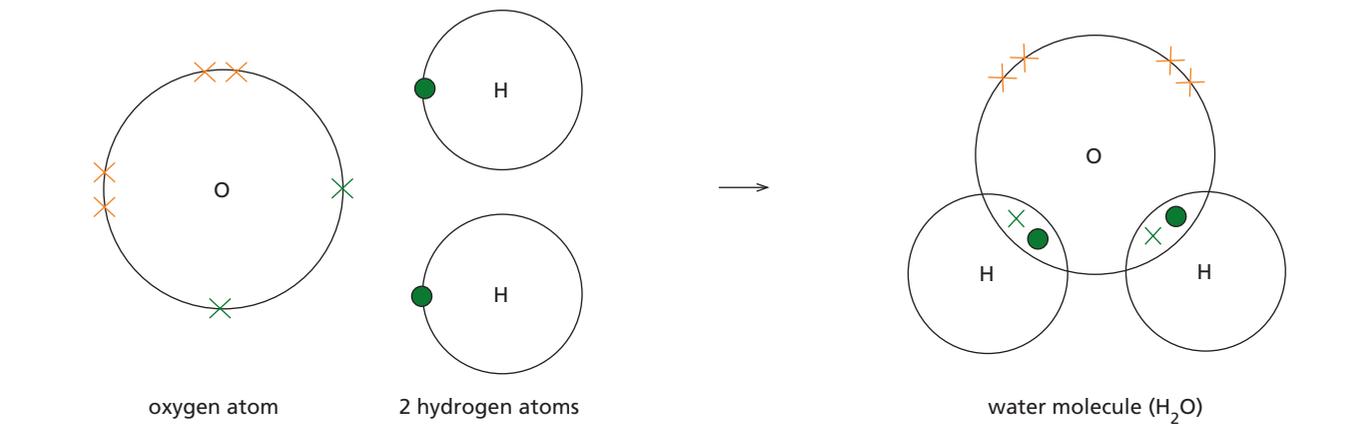
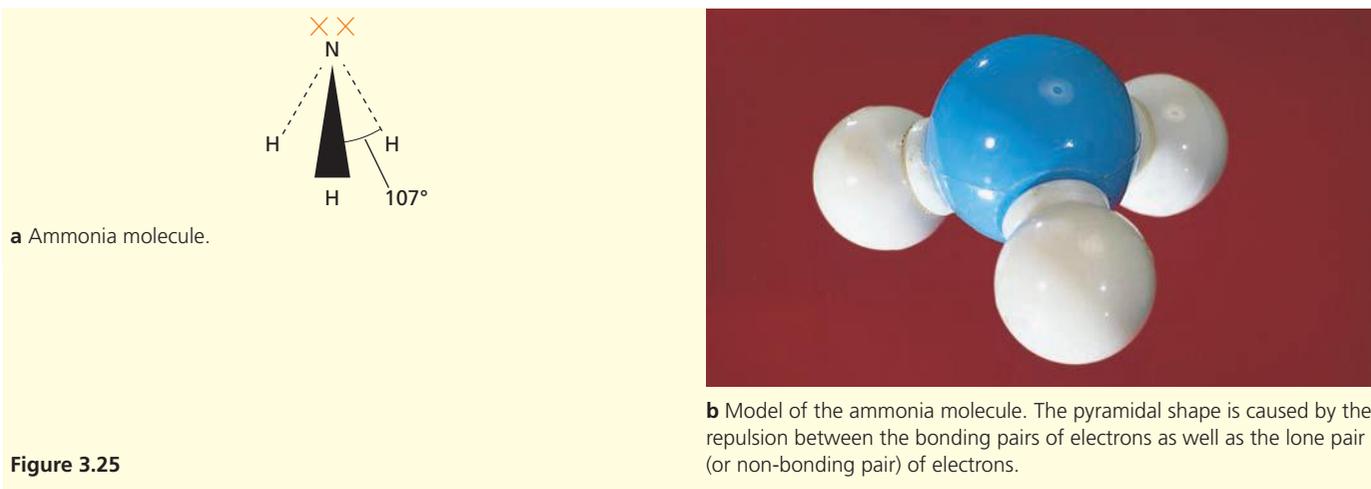
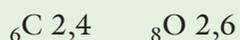


Figure 3.24 Formation of ammonia.

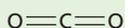
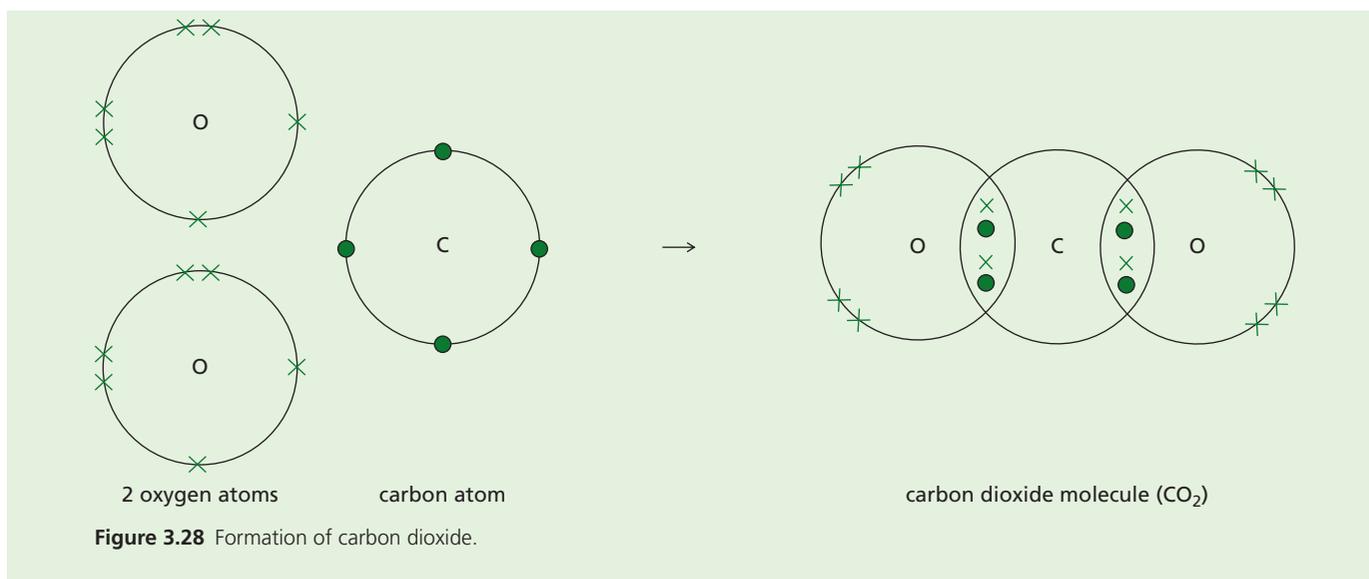


**Carbon dioxide** is a gas containing the elements carbon and oxygen. The electronic structures of the two elements are:



In this case each carbon atom needs to share four electrons to gain the electron configuration of neon.

Each oxygen needs to share two electrons to gain the electron configuration of neon. This is achieved by forming two **double covalent bonds** in which two pairs of electrons are shared in each case, as shown in Figure 3.28. Carbon dioxide is a linear molecule (Figure 3.29).



**a** Carbon dioxide molecule. Note the double covalent bond is represented by a double line.



**b** Model of the linear carbon dioxide molecule.

**Figure 3.29**

## Questions

- Draw diagrams to represent the bonding in each of the following covalent compounds:
 

<b>a</b> tetrachloromethane (CCl <sub>4</sub> )	<b>e</b> ethene (C <sub>2</sub> H <sub>4</sub> )
<b>b</b> oxygen gas (O <sub>2</sub> )	<b>f</b> methanol (CH <sub>3</sub> OH)
<b>c</b> hydrogen sulfide (H <sub>2</sub> S)	<b>g</b> nitrogen (N <sub>2</sub> ).
<b>d</b> hydrogen chloride (HCl)	
- Explain why the water molecule in Figure 3.27 is V-shaped.

## Covalent structures

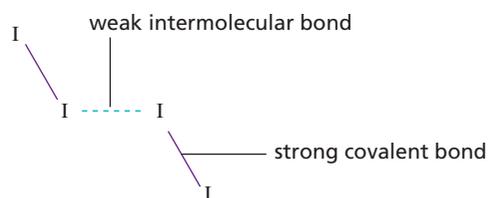
Compounds containing covalent bonds have molecules whose structures can be classified as either **simple molecular** or **giant molecular**.

Simple molecular structures are simple, formed from only a few atoms. They have strong covalent bonds between the atoms within a molecule (**intramolecular bonds**) but have weak bonds between the molecules (**intermolecular bonds**).

Some of the strongest of these weak intermolecular forces occur between water molecules.

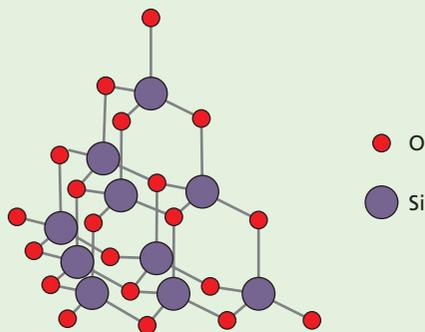
One type of weak bond between molecules is known as the **van der Waals' bond** (or force), and these forces increase steadily with the increasing size of the molecule.

Examples of simple molecules are iodine (Figure 3.30), methane, water and ethanol.



**Figure 3.30** Strong covalent and weak intermolecular forces in iodine.

**Giant molecular** or **macromolecular** structures contain many hundreds of thousands of atoms joined by strong covalent bonds. Examples of substances showing this type of structure are diamond, graphite, silicon(IV) oxide (Figure 3.31) and plastics such as polythene (Chapter 14, p. 226). Plastics are a tangled mass of very long molecules in which the atoms are joined together by strong covalent bonds to form long chains. Molten plastics can be made into fibres by being forced through hundreds of tiny holes in a ‘spinneret’ (Figure 15.19, p. 241). This process aligns the long chains of atoms along the length of the fibre. For a further discussion of plastics, see Chapter 14, pp. 226–9 and Chapter 15, p. 241.



a The silicon(IV) oxide structure in quartz.



b Quartz is a hard solid at room temperature. It has a melting point of 1610 °C and a boiling point of 2230 °C.

Figure 3.31

## Properties of covalent compounds

Covalent compounds have the following properties.

- As simple molecular substances, they are usually gases, liquids or solids with low melting and boiling points. The melting points are low because of the weak intermolecular forces of attraction which exist between simple molecules. These are weaker compared to the strong covalent bonds. Giant

molecular substances have higher melting points, because the whole structure is held together by strong covalent bonds.

It should be noted that in ionic compounds the interionic forces are much stronger than the intermolecular forces in simple covalent substances and so the melting and boiling points are generally higher.

- Generally, they do not conduct electricity when molten or dissolved in water. This is because they do not contain ions. However, some molecules actually react with water to form ions. For example, hydrogen chloride gas produces aqueous hydrogen ions and chloride ions when it dissolves in water:



- Generally, they do not dissolve in water. However, water is an excellent solvent and can interact with and dissolve some covalent molecules better than others. Covalent substances are generally soluble in organic solvents. For a further discussion of solubility of substances in organic solvents see Chapters 14 and 15.

## Allotropy

When an element can exist in more than one physical form in the same state it is said to exhibit **allotropy** (or polymorphism). Each of the different physical forms is called an **allotrope**. Allotropy is actually quite a common feature of the elements the Periodic Table (p. 136). Some examples of elements which show allotropy are sulfur, tin, iron and carbon.

## Allotropes of carbon

Carbon is a non-metallic element which exists in more than one solid structural form. Its allotropes are called **graphite** and **diamond**. Each of the allotropes has a different structure (Figures 3.32 and 3.33) and so the allotropes exhibit different physical properties (Table 3.7). The different physical properties that they exhibit lead to the allotropes being used in different ways (Table 3.8 and Figure 3.34).

Table 3.7 Physical properties of graphite and diamond.

Property	Graphite	Diamond
Appearance	A dark grey, shiny solid	A colourless transparent crystal which sparkles in light
Electrical conductivity	Conducts electricity	Does not conduct electricity
Hardness	A soft material with a slippery feel	A very hard substance
Density/g cm <sup>-3</sup>	2.25	3.51

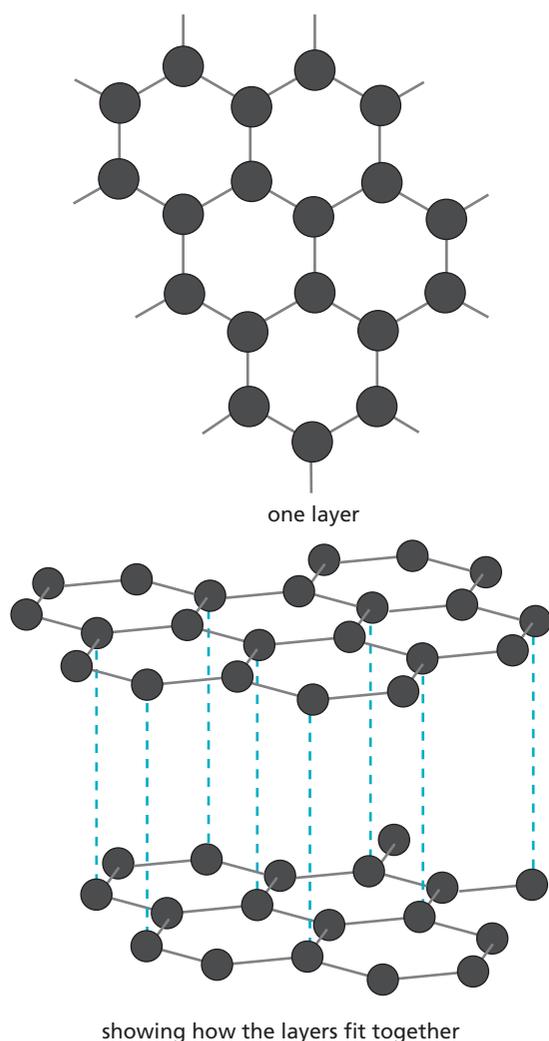
**Table 3.8** Uses of graphite and diamond.

Graphite	Diamond
Pencils	Jewellery
Electrodes	Glass cutters
Lubricant	Diamond-studded saws
	Drill bits
	Polishers

## Graphite

Figure 3.32a shows the structure of graphite. This is a layer structure. Within each layer each carbon atom is bonded to three others by strong covalent bonds. Each layer is therefore like a giant molecule. Between these layers there are weak forces of attraction (van der Waals' forces) and so the layers will pass over each other easily.

With only three covalent bonds formed between carbon atoms within the layers, an unbonded electron



a A portion of the graphite structure.

**Figure 3.32**

is present on each carbon atom. These 'spare' (or **delocalised**) electrons form electron clouds between the layers and it is because of these spare electrons that graphite conducts electricity.

In recent years a set of interesting compounds known as **graphitic compounds** have been developed. In these compounds different atoms have been fitted in between the layers of carbon atoms to produce a substance with a greater electrical conductivity than pure graphite.

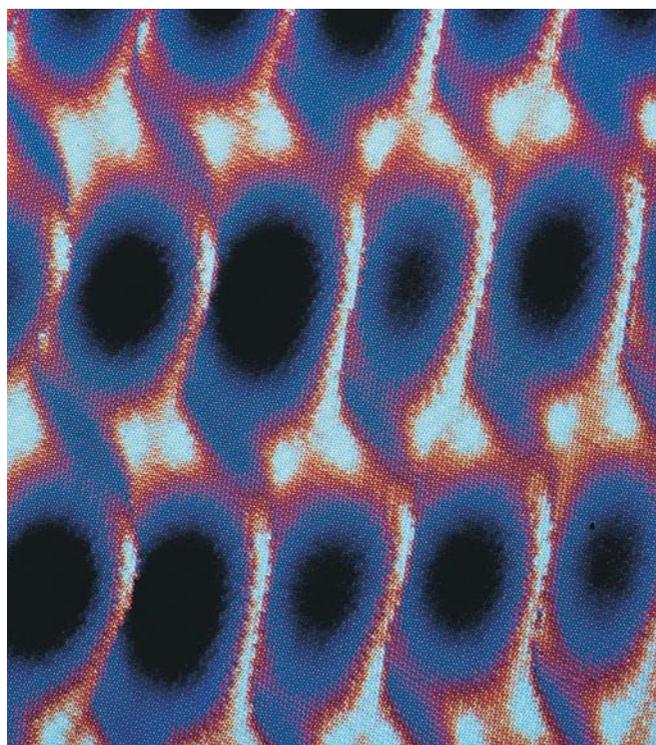
Graphite is also used as a component in certain sports equipment, such as tennis and squash rackets.

## Graphene

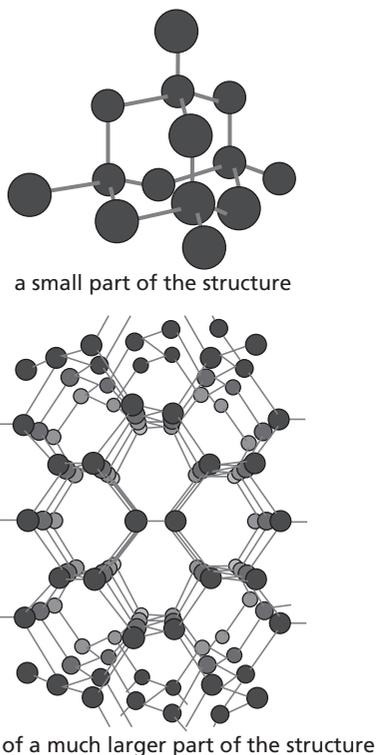
Discovered in 2004, graphene is a so-called super material made up of single layers of graphite as shown in the upper diagram of Figure 3.32a. It is able to conduct electricity one million times better than copper metal and has enormous potential in electronics.

## Diamond

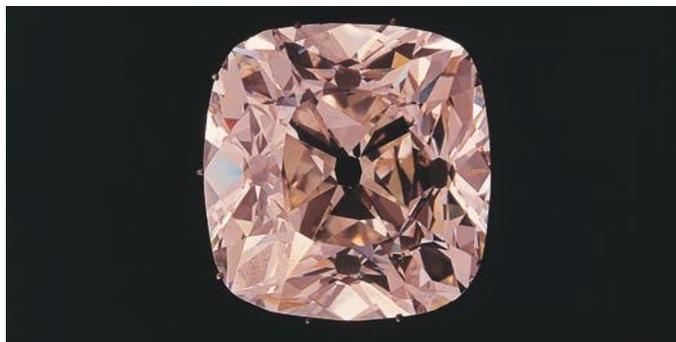
Figure 3.33a shows the diamond structure. Each of the carbon atoms in the giant structure is covalently bonded to four others. They form a tetrahedral arrangement similar to that found in silicon(IV) oxide (p. 50). This bonding scheme gives rise to a very



b A piece of graphite as imaged through a scanning tunnelling microscope.



a The structure of diamond.



b The Regent Diamond has been worn by Queen Elizabeth II.

Figure 3.33

rigid, three-dimensional structure and accounts for the extreme hardness of the substances silicon(IV) oxide and diamond. All the outer energy level electrons of the carbon atoms are used to form covalent bonds, so there are no electrons available to enable diamond to conduct electricity.

It is possible to manufacture the different allotropes of carbon. Diamond is made by heating graphite to about 300 °C at very high pressures. Diamond made by this method is known as industrial diamond. Graphite can be made by heating a mixture of coke and sand at a very high temperature in an electric arc furnace for about 24 hours.

The various uses of graphite and diamond result from their differing properties (Figure 3.34).

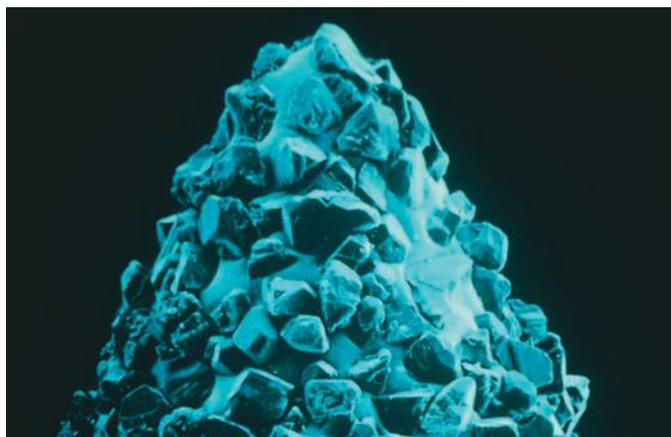
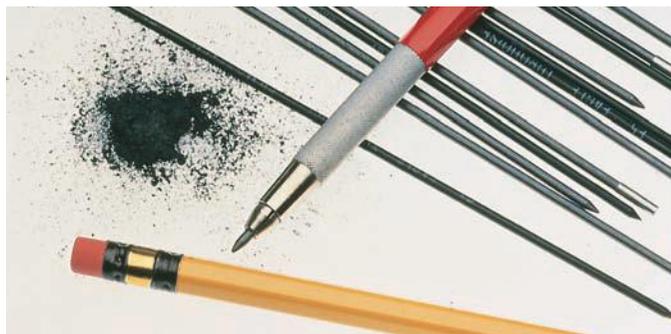
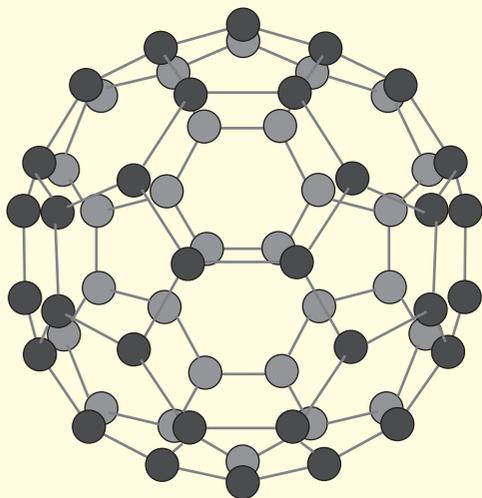


Figure 3.34 Uses of graphite (as a pencil 'lead' and in a squash racket) and diamond (as a toothed saw to cut marble and on a dentist's drill).

### Buckminsterfullerene – an unusual form of carbon

In 1985 a new allotrope of carbon was obtained by Richard Smalley and Robert Curl of Rice University, Texas. It was formed by the action of a laser beam on a sample of graphite.

The structure of buckminsterfullerene can be seen in Figure 3.35.



**Figure 3.35** Buckminsterfullerene – a ‘bucky ball’ ( $C_{60}$ ).

This spherical structure is composed of 60 carbon atoms covalently bonded together. Further spherical forms of carbon, ‘bucky balls’, containing 70, 72 and 84 carbon atoms have been identified and the discovery has led to a whole new branch of inorganic carbon chemistry. It is thought that this type of molecule exists in chimney soot. Chemists have suggested that due to the large surface area of the bucky balls they may have uses as catalysts (Chapter 7, p. 109). Also they may have uses as superconductors.

Buckminsterfullerene is named after an American architect, Buckminster Fuller, who built complex geometrical structures (Figure 3.36).

#### Questions

- 1 Explain the difference between ionic and covalent bonding. Discuss in what ways the electronic structure of the noble gases is important in both of these theories of bonding.
- 2 Compare the structures of silicon(IV) oxide (p. 50) and diamond, and describe their physical properties in terms of these structures.

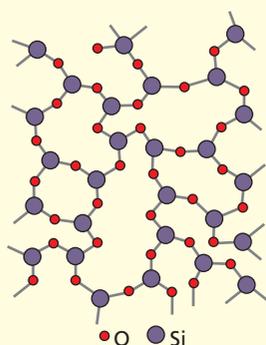


**Figure 3.36**  $C_{60}$  has a structure similar to a football and to the structure of the domes at the Eden Centre, Cornwall, UK.

## Glasses and ceramics

### Glasses

**Glasses** are irregular giant molecular structures held together by strong covalent bonds. Glass can be made by heating silicon(IV) oxide with other substances until a thick viscous liquid is formed. As this liquid cools, the atoms present cannot move freely enough to return to their arrangement within the pure silicon(IV) oxide structure. Instead they are forced to form a disordered arrangement as shown in Figure 3.37. Glass is called a **supercooled liquid**.



**Figure 3.37** Two-dimensional structure of silicon(IV) oxide.

The glass used in bottles and windows is **soda glass**. This type of glass is made by heating a mixture of sand (silicon(IV) oxide), soda (sodium carbonate) and lime (calcium oxide). Pyrex is a borosilicate glass (Figure 3.38). It is made by incorporating some boron oxide into the silicon(IV) oxide structure so that silicon atoms are replaced by boron atoms. This type of glass is tougher than soda glass and more resistant to temperature changes. It is, therefore, used in the manufacture of cooking utensils and laboratory glassware.



**Figure 3.38** This glassware is made from Pyrex, a heat-resistant glass.

### Ceramics

The word **ceramic** comes from the Greek word *keramos* meaning pottery or ‘burnt stuff’. Clay dug from the ground contains a mixture of several materials. The main one is a mineral called kaolinite,  $\text{Al}_2\text{Si}_2\text{O}_5(\text{OH})_4$ , in which the atoms are arranged in layers in a giant structure. While wet, the clay can be moulded because the kaolinite crystals move over one another. However, when it is dry the clay becomes rigid because the crystals stick together.

During firing in a furnace, the clay is heated to a temperature of  $1000^\circ\text{C}$ . A complicated series of chemical changes take place, new minerals are formed and some of the substances in the clay react to form a type of glass. The material produced at the end of the firing, the ceramic, consists of many minute mineral crystals bonded together with glass.

Modern ceramic materials now include zirconium oxide ( $\text{ZrO}_2$ ), titanium carbide (TiC), and silicon nitride (SiN). There are now many more uses of these new ceramic materials. For example, vehicle components such as ceramic bearings do not need lubrication – even at high speeds. In space technology, ceramic tiles protected the Space Shuttle from intense heat during its re-entry into the Earth’s atmosphere. In the power supply industry, they are used as insulators due to the fact that they do not conduct electricity (Figure 3.39).



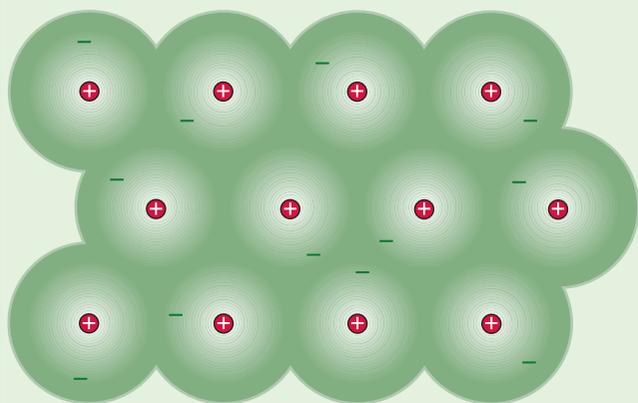
**Figure 3.39** Ceramic discs prevent short-circuits in the national grid.

## Questions

- 1 Draw up a table to summarise the properties of the different types of substances you have met in this chapter. Your table should include examples from ionic substances, covalent substances (simple and giant), ceramics and glasses.
- 2 Use your research skills, including using suitable websites, to discover details of recently developed bioceramics as well as ceramics used as superconductors.

## ● Metallic bonding

Another way in which atoms obtain a more stable electron structure is found in metals. The electrons in the outer energy level of the atom of a metal move freely throughout the structure (they are **delocalised** forming a mobile 'sea' of electrons, (Figure 3.40). When the metal atoms lose these electrons, they form a lattice of positive ions. Therefore, metals consist of positive ions embedded in moving clouds of electrons. The negatively charged electrons attract all the positive metal ions and bond them together with strong electrostatic forces of attraction as a single unit. This is the **metallic bond**.



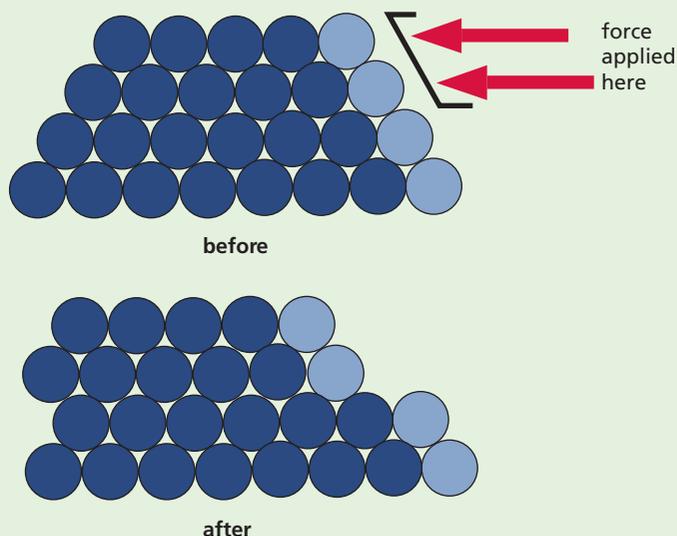
**Figure 3.40** Metals consist of positive ions surrounded by a 'sea' of electrons.

## Properties of metals

Metals have the following properties.

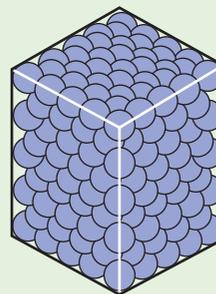
- They usually have high melting and boiling points due to the strong attraction between the positive metal ions and the mobile 'sea' of electrons.

- They conduct electricity due to the mobile electrons within the metal structure. When a metal is connected in a circuit, the electrons move towards the positive terminal while at the same time electrons are fed into the other end of the metal from the negative terminal.
- They are malleable and ductile. Unlike the fixed bonds in diamond, metallic bonds are not rigid but are still strong. If a force is applied to a metal, rows of ions can slide over one another. They reposition themselves and the strong bonds re-form as shown in Figure 3.41. Malleable means that metals can be hammered into different shapes. Ductile means that the metals can be pulled out into thin wires.



**Figure 3.41** The positions of the positive ions in a metal before and after a force has been applied.

- They have high densities because the atoms are very closely packed in a regular manner as can be seen in Figure 3.42. Different metals show different types of packing and in doing so they produce the arrangements of ions shown in Figure 3.43.



**Figure 3.42** Arrangement of ions in the crystal lattice of a metal.



- **Relative atomic mass** Symbol  $A_r$ .

$$A_r = \frac{\text{average mass of isotopes of the element}}{\frac{1}{12} \times \text{mass of 1 atom of carbon-12}}$$

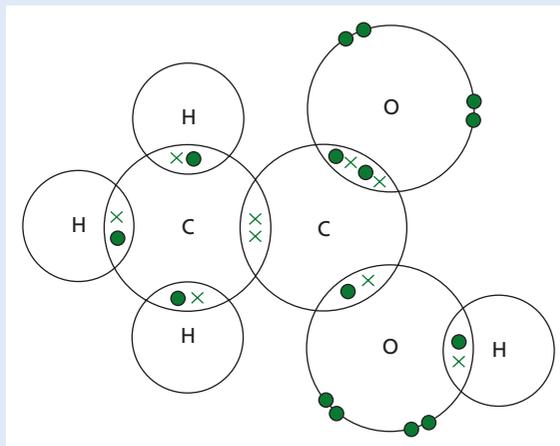
- **Simple molecular substance** A substance that possesses between one and a few hundred atoms per molecule.
- **Supercooled liquid** One which has cooled below its freezing point without solidification.
- **Valency** The combining power of an atom or group of atoms. The valency of an ion is equal to its charge. The charge is often referred to as the atom's **oxidation state**.
- **X-ray diffraction** A technique often used to study crystal structures.



# Atomic structure and bonding

## Additional questions

- 1 a Define the terms: proton, neutron and electron.  
b An atom **X** has a proton number of 19 and relative atomic mass of 39.  
(i) How many electrons, protons and neutrons are there in an atom of **X**?  
(ii) How many electrons will there be in the outer energy level (shell) of an atom of **X**?  
(iii) What is the electronic configuration of **X**?
- 2 a  $^{69}_{31}\text{Ga}$  and  $^{71}_{31}\text{Ga}$  are isotopes of gallium. With reference to this example, explain what you understand by the term isotope.  
b A sample of gallium contains 60% of atoms of  $^{69}_{31}\text{Ga}$  and 40% of atoms of  $^{71}_{31}\text{Ga}$ . Calculate the relative atomic mass of this sample of gallium.
- 3 Draw diagrams to show the bonding in each of the following compounds:  
a calcium fluoride ( $\text{CaF}_2$ )  
b oxygen ( $\text{O}_2$ )  
c magnesium chloride ( $\text{MgCl}_2$ )  
d tetrachloromethane ( $\text{CCl}_4$ ).
- 4 Use the information given in Table 3.6 on p. 43 to work out the formula for:  
a silver oxide                      e iron(II) nitrate  
b zinc chloride                    f copper(II) carbonate  
c potassium sulfate                g iron(III) hydroxide  
d calcium nitrate                 h aluminium fluoride.
- 5 Atoms of elements **X**, **Y** and **Z** have 16, 17 and 19 electrons, respectively. Atoms of argon have 18 electrons.  
a Determine the formulae of the compounds formed by the combination of the atoms of the elements:  
(i) **X** and **Z**  
(ii) **Y** and **Z**  
(iii) **X** with itself.  
b In each of the cases shown in a(i)–(iii) above, name the type of chemical bond formed.  
c Give two properties you would expect to be shown by the compounds formed in a(ii) and a(iii).
- 6 The diagram shows the arrangement of the outer electrons only in a molecule of ethanoic acid.  
a Name the different elements found in this compound.  
b What is the total number of atoms present in this molecule?



- c Between which two atoms is there a double covalent bond?  
d How many single covalent bonds does each carbon atom have?  
e Write a paragraph explaining the sorts of properties you would expect this substance to have.
- 7 The elements sodium and chlorine react together to form the compound sodium chloride, which has a giant ionic lattice structure.  
a What type of structure do the elements  
(i) sodium and (ii) chlorine have?  
b Draw a diagram to represent how the ions are arranged in the crystal lattice of sodium chloride.  
c Explain how the ions are held together in this crystal lattice.  
d Draw diagrams to show how the electrons are arranged in a sodium ion and a chloride ion. (The proton numbers of sodium and chlorine are 11 and 17, respectively.)  
e Make a table showing the properties of the three substances sodium, chlorine and sodium chloride. Include in your table:  
(i) the physical state at room temperature  
(ii) solubility in (or reaction with) water  
(iii) colour  
(iv) electrical conductivity.
- 8 Explain the following.  
a Ammonia is a gas at room temperature.  
b The melting points of sodium chloride and iodine are very different.  
c Metals generally are good conductors of electricity.  
d Buckminsterfullerene is an allotrope of carbon.  
e Graphite conducts electricity but diamond does not.

## 4

## Stoichiometry – chemical calculations

## Relative atomic mass

## Reacting masses

## Calculating moles

Moles and elements  
Moles and compounds  
Moles and gases  
Moles and solutions

## Calculating formulae

Finding the formula

## Moles and chemical equations

## Checklist

## Additional questions

## Relative atomic mass

There are at present 118 different elements known. The atoms of these elements differ in mass because of the different numbers of protons, neutrons and electrons they contain. The actual mass of one atom is very small. For example, the mass of a single atom of sulfur is around:

0.000 000 000 000 000 000 000 053 16 g

This small quantity is not easy to work with so, as you saw in Chapter 3, a scale called the **relative atomic mass** scale is used. In this scale an atom of carbon is given a relative atomic mass,  $A_r$ , of 12.00. All other atoms of the other elements are given a relative atomic mass compared to that of carbon.

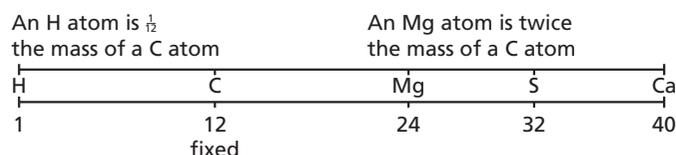
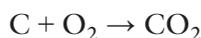


Figure 4.1 The relative atomic masses of the elements H, C, Mg, S and Ca.

## Reacting masses

Chemists often need to be able to show the relative masses of the atoms involved in a chemical process. For example, what mass of carbon dioxide would be produced if 6 g of carbon was completely combusted?



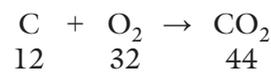
Instead of using the actual masses of atoms we use the relative atomic mass to help us answer this type of question.

In the example above we can work out the **relative formula mass** of molecules such as  $\text{O}_2$  and  $\text{CO}_2$  using the relative atomic masses of the atoms they are made from. The RFM is the sum of the relative atomic masses of all those elements shown in the formula of the substance.

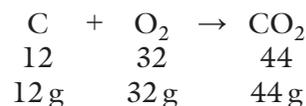
$\text{O}_2$  has a relative formula mass of  $2 \times 16 = 32$

$\text{CO}_2$  has a relative formula mass of  $12 + (2 \times 16) = 44$

So we can now use the equation to answer the question asked earlier.



Converting these relative masses to actual masses by adding mass units, g, would give:



The above calculation shows that if 12 g of carbon were burned completely then 44 g of carbon dioxide gas would be formed. So 6 g of carbon burning would result in the formation of 22 g of carbon dioxide gas.

Let us look at another example. What mass of hydrogen gas would be produced if 46 g of sodium was reacted with water? First of all write down the balanced chemical equation:



Next find the relative atomic mass of sodium from the Periodic Table (p. 136) and work out the relative formula masses of water, sodium hydroxide and hydrogen gas.

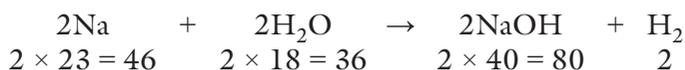
Relative atomic mass of sodium is 23.

Relative formula mass of water,  $\text{H}_2\text{O}$ , is  $(2 \times 1) + 16 = 18$

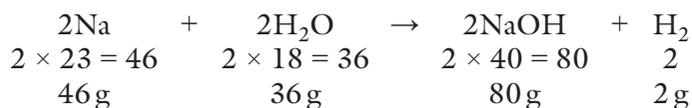
Relative formula mass of sodium hydroxide is  $23 + 16 + 1 = 40$

Relative formula mass of hydrogen gas,  $\text{H}_2$ , is  $2 \times 1 = 2$

Now you can write these masses under the balanced chemical equation taking into account the numbers used to balance the equation.



These relative masses can now be converted into actual or reacting masses by putting in mass units, for example, grams.

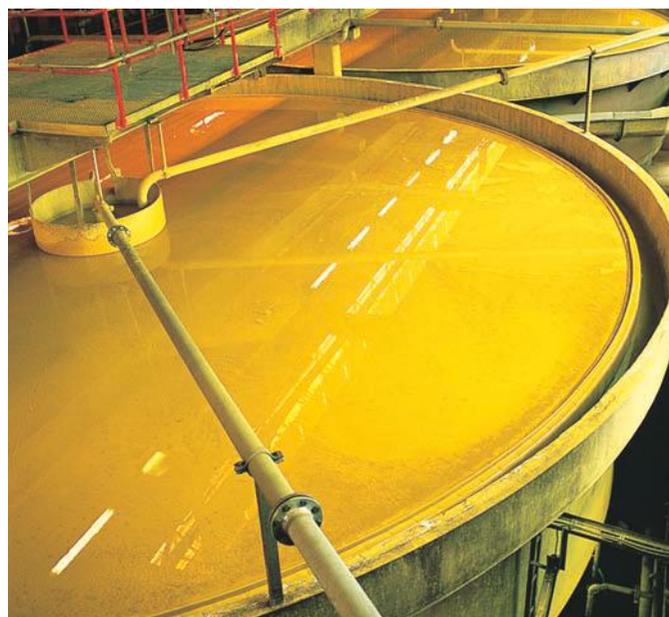


So the answer to the question of what mass of hydrogen would be produced if 46 g of sodium was reacted with water is 2 g.

### Questions

- 1 What mass of carbon dioxide gas would be produced if 10 g of calcium carbonate reacted with an excess of hydrochloric acid?
- 2 What mass of sulfur dioxide would be produced if 64 tonnes of sulfur was completely reacted with oxygen gas?

Chemists often need to know how much of a substance has been formed or used up during a chemical reaction (Figure 4.2). This is particularly important in the chemical industry, where the substances being reacted (the **reactants**) and the substances being produced (the **products**) are worth thousands of pounds. Waste costs money!

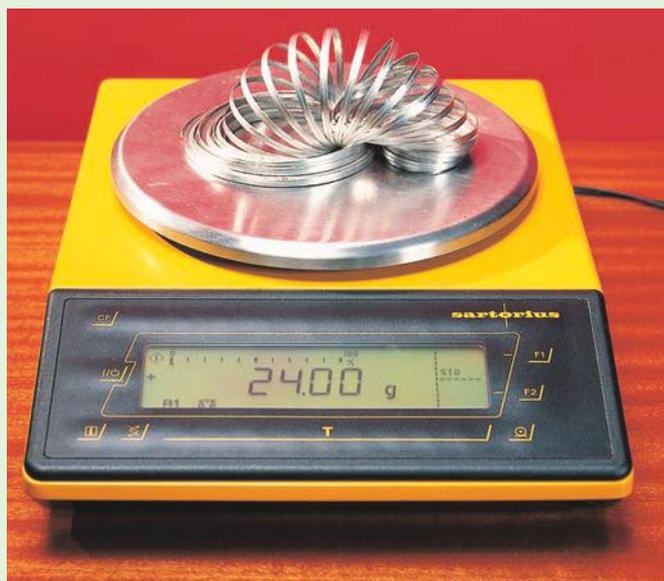


**Figure 4.2** The chemists at paint manufacturers need to know how much pigment is going to be produced.

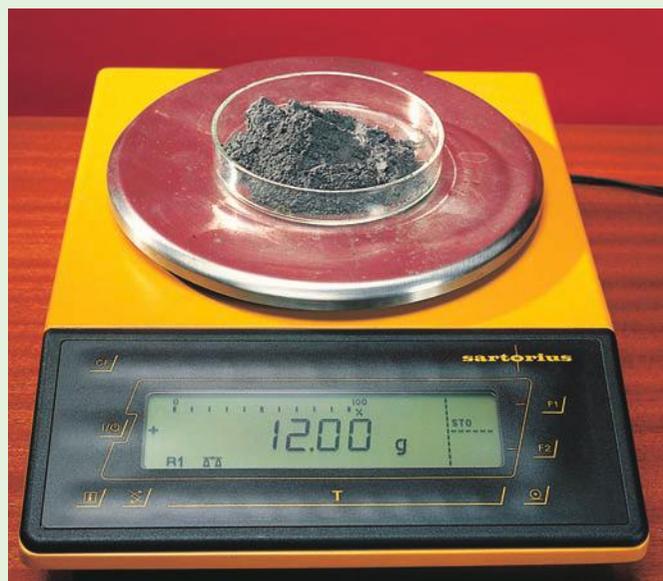
To solve this problem they need a way of counting atoms, ions or molecules. Atoms, ions and molecules are very tiny particles and it is impossible to measure out a dozen or even a hundred of them. Instead, chemists weigh out a very large number of particles.

This number is  $6 \times 10^{23}$  atoms, ions or molecules and is called **Avogadro's constant** after the famous Italian scientist Amedeo Avogadro (1776–1856). An amount of substance containing  $6 \times 10^{23}$  particles is called a **mole** (often abbreviated to mol).

So, a mole of the element magnesium is  $6 \times 10^{23}$  atoms of magnesium and a mole of the element carbon is  $6 \times 10^{23}$  atoms of carbon (Figure 4.3).



**a** A mole of magnesium.



**b** A mole of carbon.

**Figure 4.3**

## ● Calculating moles

You have already seen in Chapter 3 how we can compare the masses of all the other atoms with the mass of carbon atoms. This is the basis of the **relative atomic mass scale**. Chemists have found by experiment that if you take the relative atomic mass of an element in grams, it always contains  $6 \times 10^{23}$  or one mole of its atoms.

### Moles and elements

For example, the relative atomic mass ( $A_r$ ) of iron is 56, so one mole of iron is 56 g. Therefore, 56 g of iron contains  $6 \times 10^{23}$  atoms.

The  $A_r$  for aluminium is 27. In 27 g of aluminium it is found that there are  $6 \times 10^{23}$  atoms. Therefore, 27 g of aluminium is one mole of aluminium atoms.

The mass of a substance present in any number of moles can be calculated using the relationship:

$$\text{mass (in grams)} = \text{number of moles} \times \text{mass of 1 mole of the element}$$

#### Example 1

Calculate the mass of **a** 2 moles and **b** 0.25 mole of iron. ( $A_r$ : Fe = 56)

**a** mass of 2 moles of iron

$$= \text{number of moles} \times \text{relative atomic mass } (A_r)$$

$$= 2 \times 56$$

$$= 112 \text{ g}$$

**b** mass of 0.25 mole of iron

$$= \text{number of moles} \times \text{relative atomic mass } (A_r)$$

$$= 0.25 \times 56$$

$$= 14 \text{ g}$$

If we know the mass of the element then it is possible to calculate the number of moles of that element using:

$$\text{number of moles} = \frac{\text{mass of the element}}{\text{mass of 1 mole of that element}}$$

#### Example 2

Calculate the number of moles of aluminium present in **a** 108 g and **b** 13.5 g of the element. ( $A_r$ : Al = 27)

**a** number of moles of aluminium

$$= \frac{\text{mass of aluminium}}{\text{mass of 1 mole of aluminium}}$$

$$= \frac{108}{27}$$

$$= 4 \text{ moles}$$

**b** number of moles of aluminium

$$= \frac{\text{mass of aluminium}}{\text{mass of 1 mole of aluminium}}$$

$$= \frac{13.5}{27}$$

$$= 0.5 \text{ mole}$$

### Moles and compounds

The idea of the mole has been used so far only with elements and atoms. However, it can also be used with compounds (Figure 4.4).



**Figure 4.4** One mole of water ( $\text{H}_2\text{O}$ ) (left) and one mole of ethanol ( $\text{C}_2\text{H}_5\text{OH}$ ) (right) in separate measuring cylinders.

What is the mass of 1 mole of water ( $\text{H}_2\text{O}$ ) molecules? ( $A_r$ : H = 1; O = 16)

From the formula of water,  $\text{H}_2\text{O}$ , you will see that 1 mole of water molecules contains 2 moles of hydrogen (H) atoms and 1 mole of oxygen (O)

atoms. The mass of 1 mole of water molecules is therefore:

$$(2 \times 1) + (1 \times 16) = 18 \text{ g}$$

The mass of 1 mole of a compound is called its molar mass. If you write the molar mass of a compound without any units then it is the relative formula mass, often called the **relative molecular mass** ( $M_r$ ). So the relative formula mass of water is 18.

Now follow these examples to help you learn and understand more about moles and compounds.

### Example 1

What is **a** the mass of 1 mole and **b** the relative formula mass (RFM) of ethanol,  $\text{C}_2\text{H}_5\text{OH}$ ? ( $A_r$ : H = 1; C = 12; O = 16)

**a** One mole of  $\text{C}_2\text{H}_5\text{OH}$  contains 2 moles of carbon atoms, 6 moles of hydrogen atoms and 1 mole of oxygen atoms. Therefore:

$$\begin{aligned} \text{mass of 1 mole of ethanol} \\ &= (2 \times 12) + (6 \times 1) + (1 \times 16) \\ &= 46 \text{ g} \end{aligned}$$

**b** The RFM of ethanol is 46.

### Example 2

What is **a** the mass of 1 mole and **b** the RFM of nitrogen gas,  $\text{N}_2$ ? ( $A_r$ : N = 14)

**a** Nitrogen is a diatomic gas. Each nitrogen molecule contains two atoms of nitrogen. Therefore:

$$\begin{aligned} \text{mass of 1 mole of N}_2 \\ &= 2 \times 14 \\ &= 28 \text{ g} \end{aligned}$$

**b** The RFM of  $\text{N}_2$  is 28.

The mass of a compound found in any number of moles can be calculated using the relationship:

$$\begin{array}{l} \text{mass of} \\ \text{compound} \end{array} = \begin{array}{l} \text{number of moles} \\ \text{of the compound} \end{array} \times \begin{array}{l} \text{mass of 1 mole} \\ \text{of the compound} \end{array}$$

### Example 3

Calculate the mass of **a** 3 moles and **b** 0.2 moles of carbon dioxide gas,  $\text{CO}_2$ . ( $A_r$ : C = 12; O = 16)

**a** One mole of  $\text{CO}_2$  contains 1 mole of carbon atoms and 2 moles of oxygen atoms. Therefore:

$$\begin{aligned} \text{mass of 1 mole of CO}_2 \\ &= (1 \times 12) + (2 \times 16) \\ &= 44 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{mass of 3 moles of CO}_2 \\ &= \text{number of moles} \times \text{mass of 1 mole of CO}_2 \\ &= 3 \times 44 \\ &= 132 \text{ g} \end{aligned}$$

**b** mass of 0.2 mole of  $\text{CO}_2$   
 $= \text{number of moles} \times \text{mass of 1 mole of CO}_2$   
 $= 0.2 \times 44$   
 $= 8.8 \text{ g}$

If we know the mass of the compound then we can calculate the number of moles of the compound using the relationship:

$$\begin{array}{l} \text{number of moles} \\ \text{of compound} \end{array} = \frac{\text{mass of compound}}{\text{mass of 1 mole of the compound}}$$

### Example 4

Calculate the number of moles of magnesium oxide,  $\text{MgO}$ , in **a** 80 g and **b** 10 g of the compound. ( $A_r$ : O = 16; Mg = 24)

**a** One mole of  $\text{MgO}$  contains 1 mole of magnesium atoms and 1 mole of oxygen atoms. Therefore:

$$\begin{aligned} \text{mass of 1 mole of MgO} \\ &= (1 \times 24) + (1 \times 16) \\ &= 40 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{number of moles of MgO in 80 g} \\ &= \frac{\text{mass of MgO}}{\text{mass of 1 mole of MgO}} \\ &= \frac{80}{40} \\ &= 2 \text{ moles} \end{aligned}$$

**b** number of moles of  $\text{MgO}$  in 10 g  
 $= \frac{\text{mass of MgO}}{\text{mass of 1 mole of MgO}}$   
 $= \frac{10}{40}$   
 $= 0.25 \text{ mole}$

## Moles and gases

Many substances exist as gases. If we want to find the number of moles of a gas we can do this by measuring the volume rather than the mass.

Chemists have shown by experiment that:

One mole of any gas occupies a volume of approximately  $24 \text{ dm}^3$  (24 litres) at room temperature and pressure (rtp). This quantity is also known as the molar gas volume,  $V_m$ .

Therefore, it is relatively easy to convert volumes of gases into moles and moles of gases into volumes using the following relationship:

$$\text{number of moles of a gas} = \frac{\text{volume of the gas (in } \text{dm}^3 \text{ at rtp)}}{24 \text{ dm}^3}$$

or

$$\text{volume of a gas} = \text{number of moles of gas} \times 24 \text{ dm}^3 \text{ (in } \text{dm}^3 \text{ at rtp)}$$

### Example 1

Calculate the number of moles of ammonia gas,  $\text{NH}_3$ , in a volume of  $72 \text{ dm}^3$  of the gas measured at rtp.

$$\begin{aligned} \text{number of moles of ammonia} &= \frac{\text{volume of ammonia in } \text{dm}^3}{24 \text{ dm}^3} \\ &= \frac{72}{24} \\ &= 3 \end{aligned}$$

### Example 2

Calculate the volume of carbon dioxide gas,  $\text{CO}_2$ , occupied by **a** 5 moles and **b** 0.5 mole of the gas measured at rtp.

$$\begin{aligned} \text{a volume of } \text{CO}_2 &= \text{number of moles of } \text{CO}_2 \times 24 \text{ dm}^3 \\ &= 5 \times 24 \\ &= 120 \text{ dm}^3 \end{aligned}$$

$$\begin{aligned} \text{b volume of } \text{CO}_2 &= \text{number of moles of } \text{CO}_2 \times 24 \text{ dm}^3 \\ &= 0.5 \times 24 \\ &= 12 \text{ dm}^3 \end{aligned}$$

The volume occupied by one mole of any gas must contain  $6 \times 10^{23}$  molecules. Therefore, it follows that equal volumes of all gases measured at the same temperature and pressure must contain the

same number of molecules. This idea was also first put forward by Amedeo Avogadro and is called **Avogadro's Law**.

## Moles and solutions

Chemists often need to know the concentration of a solution. Sometimes it is measured in grams per cubic decimetre ( $\text{g dm}^{-3}$ ) but more often concentration is measured in **moles per cubic decimetre ( $\text{mol dm}^{-3}$ )**. When 1 mole of a substance is dissolved in water and the solution is made up to  $1 \text{ dm}^3$  ( $1000 \text{ cm}^3$ ), a **1 molar ( $1 \text{ mol dm}^{-3}$ )** solution is produced. Chemists do not always need to make up such large volumes of solution. A simple method of calculating the concentration is by using the relationship:

$$\text{concentration (in } \text{mol dm}^{-3}\text{)} = \frac{\text{number of moles}}{\text{volume (in } \text{dm}^3\text{)}}$$

### Example 1

Calculate the concentration (in  $\text{mol dm}^{-3}$ ) of a solution of sodium hydroxide,  $\text{NaOH}$ , which was made by dissolving 10 g of solid sodium hydroxide in water and making up to  $250 \text{ cm}^3$ . ( $A_r$ : H = 1; O = 16; Na = 23)

1 mole of  $\text{NaOH}$  contains 1 mole of sodium, 1 mole of oxygen and 1 mole of hydrogen. Therefore:

$$\begin{aligned} \text{mass of 1 mole of } \text{NaOH} &= (1 \times 23) + (1 \times 16) + (1 \times 1) \\ &= 40 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{number of moles of } \text{NaOH} \text{ in } 10 \text{ g} &= \frac{\text{mass of } \text{NaOH}}{\text{mass of 1 mole of } \text{NaOH}} \\ &= \frac{10}{40} \\ &= 0.25 \end{aligned}$$

$$(250 \text{ cm}^3 = \frac{250}{1000} \text{ dm}^3 = 0.25 \text{ dm}^3)$$

$$\begin{aligned} \text{concentration of the } \text{NaOH} \text{ solution} &= \frac{\text{number of moles of } \text{NaOH}}{\text{volume of solution (} \text{dm}^3\text{)}} \\ &= \frac{0.25}{0.25} \\ &= 1 \text{ mol dm}^{-3} \end{aligned}$$

Sometimes chemists need to know the mass of a substance that has to be dissolved to prepare a known volume of solution at a given concentration. A simple method of calculating the number of moles and so the mass of substance needed is by using the relationship:

$$\text{number of moles} = \frac{\text{concentration (in mol dm}^{-3}\text{)} \times \text{volume in solution (in dm}^3\text{)}}{1}$$

### Example 2

Calculate the mass of potassium hydroxide, KOH, that needs to be used to prepare 500 cm<sup>3</sup> of a 2 mol dm<sup>-3</sup> solution in water. (*A<sub>r</sub>*: H = 1; O = 16; K = 39)

$$\begin{aligned} \text{number of moles of KOH} &= \frac{\text{concentration of solution (mol dm}^{-3}\text{)} \times \text{volume of solution (dm}^3\text{)}}{1} \\ &= 2 \times \frac{500}{1000} \\ &= 1 \end{aligned}$$

1 mole of KOH contains 1 mole of potassium, 1 mole of oxygen and 1 mole of hydrogen. Therefore:

$$\begin{aligned} \text{mass of 1 mole of KOH} &= (1 \times 39) + (1 \times 16) + (1 \times 1) \\ &= 56 \text{ g} \end{aligned}$$

Therefore:

$$\begin{aligned} \text{mass of KOH in 1 mole} &= \text{number of moles} \times \text{mass of 1 mole} \\ &= 1 \times 56 \\ &= 56 \text{ g} \end{aligned}$$

### Questions

Use the values of *A<sub>r</sub>*, which follow to answer the questions below.

H = 1; C = 12; N = 14; O = 16; Ne = 20; Na = 23; Mg = 24; S = 32; K = 39; Fe = 56; Cu = 63.5; Zn = 65.  
One mole of any gas at rtp occupies 24 dm<sup>3</sup>.

- Calculate the number of moles in:
  - 2 g of neon atoms
  - 4 g of magnesium atoms
  - 24 g of carbon atoms.
- Calculate the mass of:
  - 0.1 mole of oxygen molecules
  - 5 moles of sulfur atoms
  - 0.25 mole of sodium atoms.

- Calculate the number of moles in:
  - 9.8 g of sulfuric acid (H<sub>2</sub>SO<sub>4</sub>)
  - 40 g of sodium hydroxide (NaOH)
  - 720 g of iron(II) oxide (FeO).
- Calculate the mass of:
  - 2 moles of zinc oxide (ZnO)
  - 0.25 mole of hydrogen sulfide (H<sub>2</sub>S)
  - 0.35 mole of copper(II) sulfate (CuSO<sub>4</sub>).
- Calculate the number of moles at rtp in:
  - 2 dm<sup>3</sup> of carbon dioxide (CO<sub>2</sub>)
  - 240 dm<sup>3</sup> of sulfur dioxide (SO<sub>2</sub>)
  - 20 cm<sup>3</sup> of carbon monoxide (CO).
- Calculate the volume of:
  - 0.3 mole of hydrogen chloride (HCl)
  - 4.4 g of carbon dioxide
  - 34 g of ammonia (NH<sub>3</sub>).
- Calculate the concentration of solutions containing:
  - 0.2 mole of sodium hydroxide dissolved in water and made up to 100 cm<sup>3</sup>
  - 9.8 g of sulfuric acid dissolved in water and made up to 500 cm<sup>3</sup>.
- Calculate the mass of:
  - copper(II) sulfate (CuSO<sub>4</sub>) which needs to be used to prepare 500 cm<sup>3</sup> of a 0.1 mol dm<sup>-3</sup> solution
  - potassium nitrate (KNO<sub>3</sub>) which needs to be used to prepare 200 cm<sup>3</sup> of a 2 mol dm<sup>-3</sup> solution.

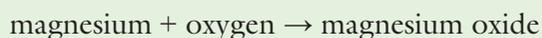
## Calculating formulae

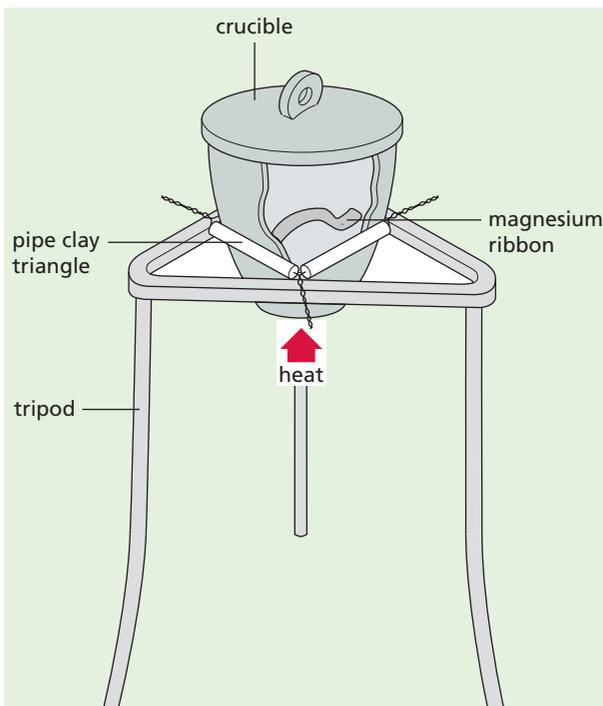
If we have 1 mole of a compound, then the formula shows the number of moles of each element in that compound. For example, the formula for lead(II) bromide is PbBr<sub>2</sub>. This means that 1 mole of lead(II) bromide contains 1 mole of lead ions and 2 moles of bromide ions. If we do not know the formula of a compound, we can find the masses of the elements present experimentally and these masses can be used to work out the formula of that compound.

### Finding the formula

#### Magnesium oxide

When magnesium ribbon is heated strongly, it burns very brightly to form the white powder called magnesium oxide (Figure 2.7, p. 14).





**Figure 4.5** Apparatus used to determine magnesium oxide's formula.

**Table 4.1** Data from the experiment shown in Figure 4.5.

Mass of crucible	14.63 g
Mass of crucible and magnesium	14.87 g
Mass of crucible and magnesium oxide	15.03 g
Mass of magnesium used	0.24 g
Mass of oxygen which has reacted with the magnesium	0.16 g

The data shown in Table 4.1 were obtained from an experiment using the apparatus shown in Figure 4.5 to find the formula for this white powder, magnesium oxide. From these data we can calculate the number of moles of each of the reacting elements. ( $A_r$ : O = 16; Mg = 24)

	<b>Mg</b>	<b>O</b>
Masses reacting (g)	0.24	0.16
Number of moles	$\frac{0.24}{24}$	$\frac{0.16}{16}$
	= 0.01	= 0.01
Ratio of moles	1	1
Formula	MgO	

This formula is the **empirical formula** of the compound. It shows the simplest ratio of the atoms present.

### Unknown compound 1

In another experiment an unknown organic compound was found to contain 0.12 g of carbon and 0.02 g of hydrogen. Calculate the empirical formula of the compound. ( $A_r$ : H = 1; C = 12)

	<b>C</b>	<b>H</b>
Masses (g)	0.12	0.02
Number of moles	$\frac{0.12}{12}$	$\frac{0.02}{1}$
	= 0.01	= 0.02
Ratio of moles	1	2
Empirical formula	CH <sub>2</sub>	

From our knowledge of bonding (Chapter 3, p. 46) we know that a molecule of this formula cannot exist. However, molecules with the following formulae do exist: C<sub>2</sub>H<sub>4</sub>, C<sub>3</sub>H<sub>6</sub>, C<sub>4</sub>H<sub>8</sub> and C<sub>5</sub>H<sub>10</sub>. All of these formulae show the same ratio of carbon atoms to hydrogen atoms, CH<sub>2</sub>, as our unknown. To find out which of these formulae is the actual formula for the unknown organic compound, we need to know the mass of one mole of the compound.

Using a mass spectrometer, the relative molecular mass ( $M_r$ ) of this organic compound was found to be 56. We need to find out the number of empirical formulae units present:

$$\begin{aligned} M_r \text{ of the empirical formula unit} \\ &= (1 \times 12) + (2 \times 1) \\ &= 14 \end{aligned}$$

Number of empirical formula units present

$$\begin{aligned} &= \frac{M_r \text{ of compound}}{M_r \text{ of empirical formula unit}} \\ &= \frac{56}{14} \\ &= 4 \end{aligned}$$

Therefore, the actual formula of the unknown organic compound is  $4 \times \text{CH}_2 = \text{C}_4\text{H}_8$ .

This substance is called butene. C<sub>4</sub>H<sub>8</sub> is the **molecular formula** for this substance and shows the **actual** numbers of atoms of each element present in one molecule of the substance.

Sometimes the composition of a compound is given as a percentage by mass of the elements present. In cases such as this the procedure shown in the next example is followed.

**Unknown compound 2**

Calculate the empirical formula of an organic compound containing 92.3% carbon and 7.7% hydrogen by mass. The  $M_r$  of the organic compound is 78. What is its molecular formula? ( $A_r$ : H = 1; C = 12)

	<b>C</b>	<b>H</b>
% by mass	92.3	7.7
Masses in 100 g	92.3 g	7.7 g
Number of moles	$\frac{92.3}{12}$	$\frac{7.7}{1}$
	= 7.7	= 7.7
Ratio of moles	1	1

Empirical formula CH  
 $M_r$  of the empirical formula unit CH  
 = 12 + 1  
 = 13

Number of empirical formula units present

$$= \frac{M_r \text{ of compound}}{M_r \text{ of empirical formula unit}}$$

$$= \frac{78}{13}$$

$$= 6$$

The molecular formula of the organic compound is  $6 \times \text{CH} = \text{C}_6\text{H}_6$ . This is a substance called benzene.

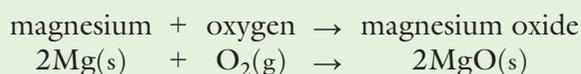
**Questions**

Use the following values of  $A_r$  to answer the questions below: H = 1; C = 12; O = 16; Ca = 40.

- Determine the empirical formula of an oxide of calcium formed when 0.4g of calcium reacts with 0.16g of oxygen.
- Determine the empirical formula of an organic hydrocarbon compound which contains 80% by mass of carbon and 20% by mass of hydrogen. If the  $M_r$  of the compound is 30, what is its molecular formula?

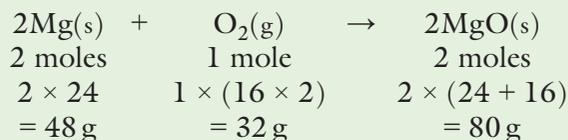
## ● Moles and chemical equations

When we write a balanced chemical equation we are indicating the numbers of moles of reactants and products involved in the chemical reaction. Consider the reaction between magnesium and oxygen.



This shows that 2 moles of magnesium react with 1 mole of oxygen to give 2 moles of magnesium oxide.

Using the ideas of moles and masses we can use this information to calculate the quantities of the different chemicals involved.



You will notice that the total mass of reactants is equal to the total mass of product. This is true for any chemical reaction and it is known as the **Law of conservation of mass**. This law was understood by the Greeks but was first clearly formulated by Antoine Lavoisier in 1774. Chemists can use this idea to calculate masses of products formed and reactants used in chemical processes before they are carried out.

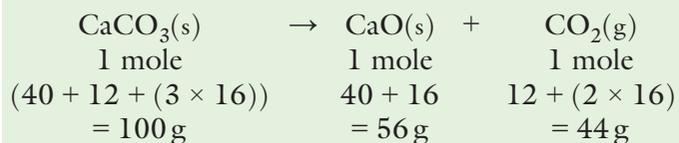
**Example 1 – using a solid**

Lime (calcium oxide, CaO) is used in the manufacture of mortar. It is manufactured in large quantities in Europe (see Figure 4.6) by heating limestone (calcium carbonate, CaCO<sub>3</sub>).



**Figure 4.6** A rotary kiln for burning (calcining) limestone into lime, located in Moha, Belgium.

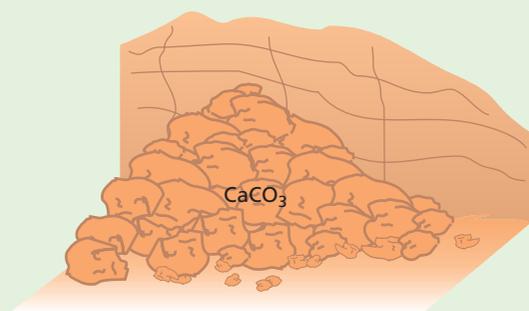
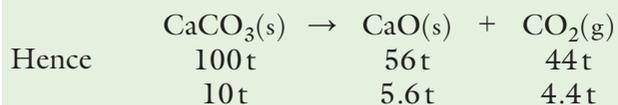
The equation for the process is:



Calculate the amount of lime produced when 10 tonnes of limestone are heated (Figure 4.7). ( $A_r$ : C = 12; O = 16; Ca = 40)

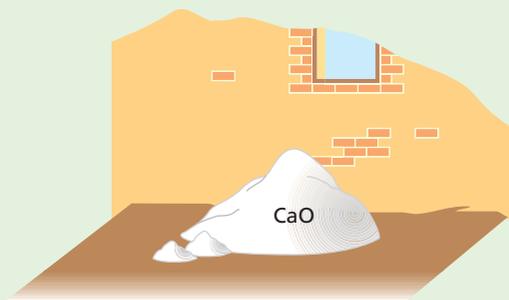
$$\begin{array}{l}
 1 \text{ tonne (t)} = 1000 \text{ kg} \\
 1 \text{ kg} = 1000 \text{ g}
 \end{array}$$

From this relationship between grams and tonnes we can replace the masses in grams by masses in tonnes.



10 tonnes of limestone

heat  
↓



? tonnes of lime

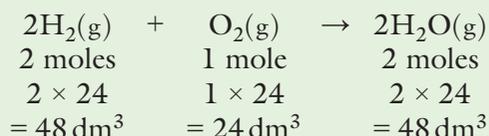
Figure 4.7 How much lime is produced?

The equation now shows that 100 t of limestone will produce 56 t of lime. Therefore, 10 t of limestone will produce 5.6 t of lime.

### Example 2 – using a gas

Many chemical processes involve gases. The volume of a gas is measured more easily than its mass. This example shows how chemists work out the volumes of gaseous reactants and products needed using Avogadro's Law and the idea of moles.

Some rockets use hydrogen gas as a fuel. When hydrogen burns in oxygen it forms water vapour. Calculate the volumes of **a**  $\text{O}_2(\text{g})$  used and **b** water,  $\text{H}_2\text{O}(\text{g})$ , produced if  $960 \text{ dm}^3$  of hydrogen gas,  $\text{H}_2(\text{g})$ , were burned in oxygen. ( $A_r$ : H = 1; O = 16) Assume 1 mole of any gas occupies a volume of  $24 \text{ dm}^3$ .



Therefore:



When  $960 \text{ dm}^3$  of hydrogen are burned in oxygen:

- a**  $480 \text{ dm}^3$  of oxygen are required and  
**b**  $960 \text{ dm}^3$  of  $\text{H}_2\text{O}(\text{g})$  are produced.

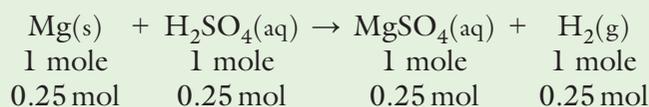
### Example 3 – using a solution

Chemists usually carry out reactions using solutions. If they know the concentration of the solution(s) they are using they can find out the quantities reacting.

Calculate the volume of  $1 \text{ mol dm}^{-3}$  solution of  $\text{H}_2\text{SO}_4$  required to react completely with 6 g of magnesium. ( $A_r$ : Mg = 24)

number of moles of magnesium

$$\begin{aligned}
 &= \frac{\text{mass of magnesium}}{\text{mass of 1 mole of magnesium}} \\
 &= \frac{6}{24} \\
 &= 0.25
 \end{aligned}$$



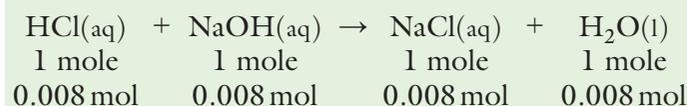
So 0.25 mol of  $\text{H}_2\text{SO}_4(\text{aq})$  is required. Using:

$$\begin{aligned} & \text{volume of } \text{H}_2\text{SO}_4(\text{aq}) \text{ (dm}^3\text{)} \\ &= \frac{\text{moles of } \text{H}_2\text{SO}_4}{\text{concentration of } \text{H}_2\text{SO}_4 \text{ (mol dm}^{-3}\text{)}} \\ &= \frac{0.25}{1} \\ &= 0.25 \text{ dm}^3 \text{ or } 250 \text{ cm}^3 \end{aligned}$$

#### Example 4 – using a solution

What is the concentration of sodium hydroxide solution used in the following neutralisation reaction?  $40 \text{ cm}^3$  of  $0.2 \text{ mol dm}^{-3}$  solution of hydrochloric acid just neutralised  $20 \text{ cm}^3$  of sodium hydroxide solution.

$$\begin{aligned} & \text{number of moles of HCl used} \\ &= \text{concentration (mol dm}^{-3}\text{)} \times \text{volume (dm}^3\text{)} \\ &= 0.2 \times 0.04 \\ &= 0.008 \end{aligned}$$



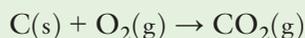
You will see that 0.008 mole of NaOH was present. The concentration of the  $\text{NaOH(aq)}$  is given by:

$$\begin{aligned} & \text{concentration of NaOH (mol dm}^{-3}\text{)} \\ &= \frac{\text{number of moles of NaOH}}{\text{volume of NaOH (dm}^3\text{)}} \\ & \left( \text{volume of NaOH in dm}^3 = \frac{20}{1000} = 0.02 \right) \\ &= \frac{0.008}{0.02} \\ &= 0.4 \text{ mol dm}^{-3} \end{aligned}$$

#### Percentage yield

Chemical reactions rarely produce the predicted amount of product from the masses of reactants in the reaction; they are not 100% efficient.

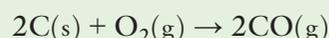
An example of this is the reaction of carbon with oxygen to produce carbon dioxide gas.



From the equation for the reaction you can see that one mole of carbon reacts with oxygen to give one mole of carbon dioxide gas.

If you burn 12 grams, 1 mole, of carbon to make  $\text{CO}_2$ , then the amount of carbon dioxide expected is 44 grams, 1 mole of  $\text{CO}_2$ . The **theoretical yield** of carbon dioxide from this reaction is 44 grams. This only occurs, however, if the reaction is 100% efficient.

The mass of carbon dioxide you will get will be less than 44 grams because another reaction can also occur between carbon and oxygen. Some carbon reacts to make carbon monoxide, CO.



The percentage yield of the reaction is based on the amount of carbon dioxide that is actually produced against what should have been produced, if the reaction was 100% efficient.

If 12 grams of carbon was burned in excess oxygen and only 28 grams of carbon dioxide was produced the percentage yield can be worked out as shown below.

$$\begin{aligned} \text{percentage yield of carbon dioxide} &= \frac{28}{44} \times 100 \\ &= 63.6\% \end{aligned}$$

In general:

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

#### Percentage purity

You saw in Chapter 2 p. 24 that the **purity** of a substance is very important.

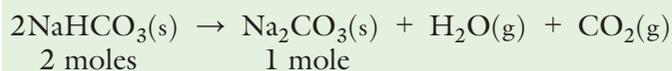
If a factory makes medicines or chemicals used in food then the purity of the product will be crucial. This is because the impurities may harm the people using the medicine or eating the food.

$$\begin{aligned} & \text{percentage purity} \\ &= \frac{\text{mass of the pure product}}{\text{mass of the impure product obtained}} \times 100\% \end{aligned}$$

Sodium hydrogencarbonate,  $\text{NaHCO}_3$ , is used in the manufacture of some toothpastes and as a raising agent in food production. The purity of this substance can be obtained by measuring how much carbon dioxide is given off.

84 grams of sodium hydrogencarbonate was thermally decomposed and  $11.5 \text{ dm}^3$  of carbon dioxide gas was collected at room temperature and pressure (rtp).

The equation for the reaction is:



Step 1: Calculate the RFM of sodium hydrogencarbonate ( $A_r$ : Na = 23; C = 12; O = 16; H = 1)  
RFM of  $\text{NaHCO}_3 = 84$

Step 2: 2 moles of  $\text{NaHCO}_3$  produces 1 mole of  $\text{CO}_2$ .  
168 g of  $\text{NaHCO}_3$  would give 44 g of  $\text{CO}_2$ , which would have a volume of  $24 \text{ dm}^3$  at rtp.  
84 g of  $\text{NaHCO}_3$  should give  $12 \text{ dm}^3$  of  $\text{CO}_2$  at rtp.  
Hence the mass of  $\text{NaHCO}_3$  in the sample was  $84 \times \frac{11.5}{12} = 80.5 \text{ g}$

Step 3: Calculate the percentage purity.  
There is 80.5 g of sodium hydrogencarbonate in the 84 g sample.  
Percentage purity =  $\frac{80.5}{84} \times 100\% = 95.8\%$

## Questions

Use the following  $A_r$  values to answer the questions below: O = 16; Mg = 24; S = 32; K = 39; Cu = 63.5.

- 1 Calculate the mass of sulfur dioxide produced by burning 16 g of sulfur in an excess of oxygen in the Contact process (see p. 199).
- 2 Calculate the mass of sulfur which, when burned in excess oxygen, produces 640 g of sulfur dioxide in the Contact process.
- 3 Calculate the mass of copper required to produce 159 g of copper(II) oxide when heated in excess oxygen.
- 4 In the rocket mentioned previously in which hydrogen is used as a fuel, calculate the volume of hydrogen used to produce  $24 \text{ dm}^3$  of water ( $\text{H}_2\text{O}(\text{g})$ ).
- 5 Calculate the volume of  $2 \text{ mol dm}^{-3}$  solution of sulfuric acid required to react with 24 g of magnesium.
- 6 What is the concentration of potassium hydroxide solution used in the following neutralisation reaction?  $20 \text{ cm}^3$  of  $0.2 \text{ mol dm}^{-3}$  solution of hydrochloric acid just neutralised  $15 \text{ cm}^3$  of potassium hydroxide solution.

## Checklist

After studying Chapter 4 you should know and understand the following terms.

- **Avogadro's Law** Equal volumes of all gases measured under the same conditions of temperature and pressure contain equal numbers of molecules.

- **Calculating moles of compounds**

$$\text{mass of compound (in grams)} = \text{number of moles} \times \text{mass of 1 mole of compound}$$

$$\text{number of moles} = \frac{\text{mass of compound}}{\text{mass of 1 mole of compound}}$$

- **Calculating moles of elements**

$$\text{mass of element (in grams)} = \text{number of moles} \times \text{mass of 1 mole of the element}$$

$$\text{number of moles} = \frac{\text{mass of the element}}{\text{mass of 1 mole of that element}}$$

- **Calculating moles of gases**

1 mole of any gas occupies  $24 \text{ dm}^3$  (litres) at room temperature and pressure (rtp).

$$\text{number of moles of gas} = \frac{\text{volume of the gas (in } \text{dm}^3 \text{ at rtp)}}{24 \text{ dm}^3}$$

- **Calculating moles of solutions**

$$\text{concentration of a solution (in } \text{mol dm}^{-3}) = \frac{\text{number of moles of solute}}{\text{volume (in } \text{dm}^3)}$$

$$\text{number of moles} = \text{concentration (in } \text{mol dm}^{-3}) \times \text{volume of solution (in } \text{dm}^3)$$

- **Empirical formula** A formula showing the simplest ratio of atoms present.
- **Mole** The amount of substance which contains  $6 \times 10^{23}$  atoms, ions or molecules. This number is called Avogadro's constant.
  - Atoms – 1 mole of atoms has a mass equal to the relative atomic mass ( $A_r$ ) in grams.
  - Molecules – 1 mole of molecules has a mass equal to the relative molecular mass ( $M_r$ ) in grams.
- **Molecular formula** A formula showing the actual number of atoms of each element present in one molecule.

- **Percentage yield**

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

- **Relative formula mass (RFM)** The sum of the relative atomic masses of all those elements shown in the formula of the substance. This is often referred to as the **relative molecular mass** ( $M_r$ ).



# Stoichiometry – chemical calculations

## Additional questions

Use the data in the table below to answer the questions which follow.

Element	$A_r$
H	1
C	12
N	14
O	16
Na	23
Mg	24
Si	28
S	32
Cl	35.5
Fe	56

- Calculate the mass of:
  - 1 mole of:
    - chlorine molecules
    - iron(III) oxide.
  - 0.5 mole of:
    - magnesium nitrate
    - ammonia.
- Calculate the volume occupied, at rtp, by the following gases. (One mole of any gas occupies a volume of  $24 \text{ dm}^3$  at rtp.)
  - 12.5 moles of sulfur dioxide gas.
  - 0.15 mole of nitrogen gas.
- Calculate the number of moles of gas present in the following:
  - $36 \text{ cm}^3$  of sulfur dioxide
  - $144 \text{ dm}^3$  of hydrogen sulfide.
- Use the following experimental information to determine the empirical formula of an oxide of silicon.

Mass of crucible	18.20 g
Mass of crucible + silicon	18.48 g
Mass of crucible + oxide of silicon	18.80 g
- Calculate the empirical formula of an organic liquid containing 26.67% of carbon and 2.22% of hydrogen, with the rest being oxygen.
  - The  $M_r$  of the liquid is 90. What is its molecular formula?
- Iron is extracted from its ore, haematite, in the blast furnace. The main extraction reaction is:
$$\text{Fe}_2\text{O}_3(\text{s}) + 3\text{CO}(\text{g}) \rightarrow 2\text{Fe}(\text{s}) + 3\text{CO}_2(\text{g})$$
  - Name the reducing agent in this process.
  - Name the oxide of iron shown in the equation.
  - Explain why this is a **redox** reaction.
  - Calculate the mass of iron which will be produced from 640 tonnes of haematite.
- Consider the following information about the newly discovered element, vulcium, whose symbol is Vu. 'Vulcium is a solid at room temperature. It is easily cut by a penknife to reveal a shiny surface which tarnishes quite rapidly. It reacts violently with water, liberating a flammable gas and forms a solution with a pH of 13. When vulcium reacts with chlorine, it forms a white crystalline solid containing 29.5% chlorine.'  
( $A_r$ : Vu = 85)
  - Calculate the empirical formula of vulcium chloride.
  - To which group of the Periodic Table (p. 136) should vulcium be assigned?
  - Write a word and balanced chemical equation for the reaction between vulcium and chlorine.
  - What other information in the description supports the assignment of group you have given to vulcium?
  - What type of bonding is present in vulcium chloride?
  - Write a word and balanced chemical equation for the reaction between vulcium and water.
  - Write the formulae for:
    - vulcium sulfate
    - vulcium carbonate
    - vulcium hydroxide.Look at the Periodic Table (p. 294) to find out the real name of vulcium.
- 0.048 g of magnesium was reacted with excess dilute hydrochloric acid at room temperature and pressure. The hydrogen gas given off was collected.
  - Write a word and balanced symbol equation for the reaction taking place.
  - Draw a diagram of an apparatus which could be used to carry out this experiment and collect the hydrogen gas.
  - How many moles of magnesium were used?
  - Using the equation you have written in your answer to a, calculate the number of moles of hydrogen and hence the volume of this gas produced.
  - Calculate the volume of a solution containing  $0.1 \text{ mol dm}^{-3}$  hydrochloric acid which would be needed to react exactly with 0.048 g of magnesium.

## 5

## Electricity and chemistry

**Electrolysis of lead(II) bromide****Electrolysis of aluminium oxide**

Anodising

**Electrolysis of aqueous solutions**

Electrolysis of dilute sulfuric acid

The chlor-alkali industry

Tests for aqueous cations

**Electrolysis of concentrated hydrochloric acid****Electrolysis of copper(II) sulfate solution**

Purification of copper

**Electrolysis guidelines****Electroplating**

Plating plastics

**Checklist****Additional questions**

What do all the items in the photographs shown in Figure 5.1 have in common? They all involve electricity through a process known as **electrolysis**. Electrolysis is the breakdown of an ionic compound, molten or in solution, by the passage of electricity through it. The substance which is decomposed is called the **electrolyte** (Figure 5.2). An electrolyte is a substance that conducts electricity when in the molten state or in solution.



**a** This watch has a thin coating of gold over steel; the thin coating is produced by electrolysis.



**b** This picture frame has been silver plated using an electroplating process involving electrolysis.



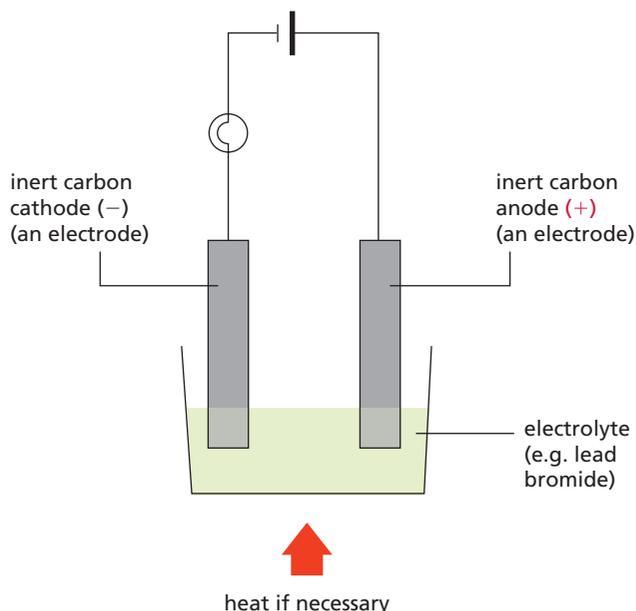
**c** Aluminium is produced by electrolysis.

**Figure 5.1**

The electricity is carried through the electrolyte by **ions**. In the molten state and in solution the ions are free to move to the appropriate electrodes due to weakened forces of attraction between them.

- Substances that do not conduct electricity when in the molten state or in solution are called **non-electrolytes**.
- Substances that conduct electricity **to a small extent** in the molten state or in solution are called **weak electrolytes**.

The electric current enters and leaves the electrolyte through **electrodes**, which are usually made of unreactive metals such as platinum or of the non-metal carbon (graphite). These are said to be **inert** electrodes because they do not react with the products of electrolysis. The names given to the two electrodes are **cathode**, the negative electrode which attracts **cations** (positively charged ions), and **anode**, the positive electrode which attracts **anions** (negatively charged ions).



**Figure 5.2** The important terms used in electrolysis.

The transfer of charge during electrolysis is by:

- the movement of electrons in the metallic or graphite electrodes
- the removal or addition of electrons from the external circuit at the electrodes
- the movement of ions in the electrolyte.

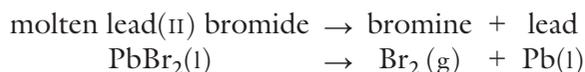
Note that the conduction which takes place in the electrodes is due to the movement of delocalised electrons (pp. 51 and 55) whereas in the electrolyte, as stated earlier, the charge carriers are ions.

Electrolysis is very important in industry. To help you to understand what is happening in the process shown in the photographs, we will first consider the electrolysis of lead(II) bromide.

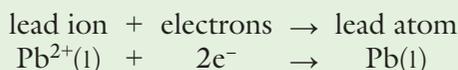
## ● Electrolysis of lead(II) bromide

Consider Figure 5.3, which shows solid lead(II) bromide ( $\text{PbBr}_2$ ) in a crucible with two carbon electrodes in contact with it. When the electrodes are first connected, the bulb does not light, because the solid compound does not allow electricity to pass through it. However, when the compound is heated until it is molten, the bulb does light. The lead(II) bromide is now behaving as an electrolyte. When this happens an orange-red vapour is seen at the anode and lead metal is produced at the cathode.

The break-up (decomposition) of lead(II) bromide into its constituent elements by the passage of an electric current is called **electrolysis**.



For lead metal to be formed, or deposited, at the cathode, the lead ions must be attracted to and move towards the electrode (Figure 5.4). To produce lead metal atoms these lead ions must each collect two electrons at the cathode:



As you saw in Chapter 3 (p. 39), this process of gaining electrons is called **reduction**.

To form bromine molecules each bromide ion must first of all move to the anode and lose its extra negative charge at the anode and so form a neutral bromine atom:

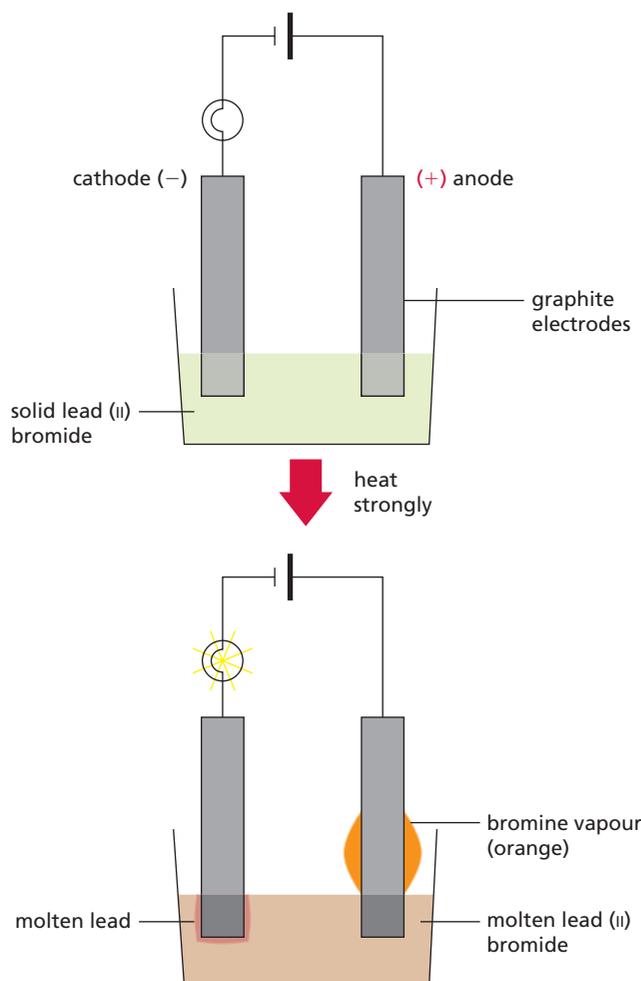


Figure 5.3 The electrolysis of molten lead(II) bromide.

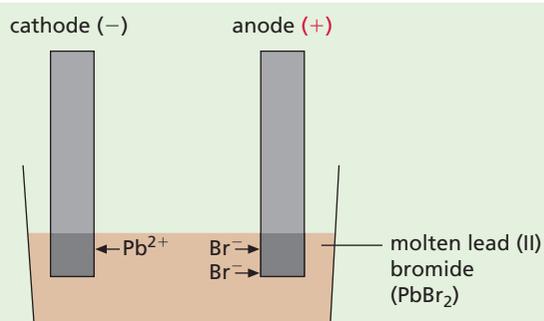
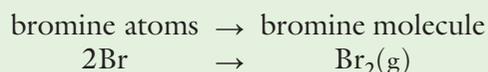


Figure 5.4 The lead ions ( $\text{Pb}^{2+}$ ) are attracted to the cathode and the bromide ions ( $\text{Br}^-$ ) are attracted to the anode.



Two bromine atoms then combine to form a bromine molecule:



As you saw in Chapter 3 (p. 39), this process of losing electrons is called **oxidation**.

**Question**

- 1 Predict the products of the electrolysis of:
- a potassium chloride
  - b lead oxide.

## ● Electrolysis of aluminium oxide

Aluminium is the most abundant metallic element in the Earth's crust. It was discovered in 1825 by Hans Christian Oersted in Copenhagen, Denmark, but was first isolated by Friedrich Wöhler in 1827. It makes up 8% of the crust and is found in the minerals bauxite (Figure 5.5), cryolite and mica, as well as in clay.

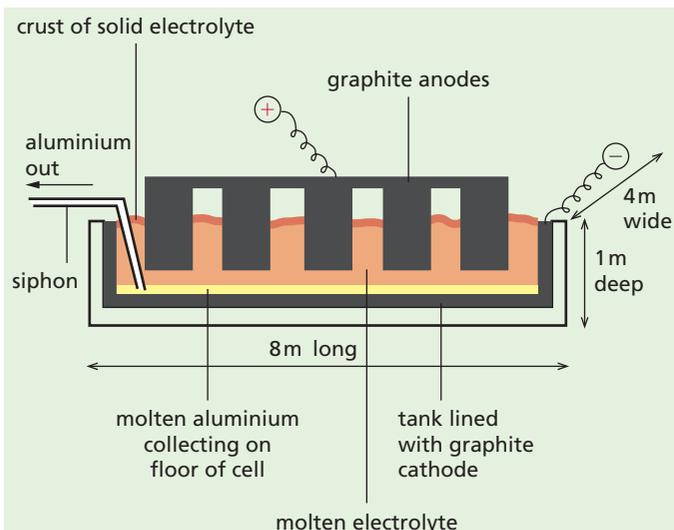
In the nineteenth century Napoleon III owned a very precious dinner service. It was said to be made of a metal more precious than gold. That metal was aluminium. The reason it was precious was that it was very rarely found as the pure metal. Aluminium is a reactive metal and as such was very difficult to extract from its ore. Reactive metals hold on tightly to the element(s) they have combined with and many are extracted from their ores by electrolysis.

Today we use aluminium in very large quantities. The annual production in the world is 19.5 million tonnes. The commercial extraction of aluminium has been made possible by two scientists, working independently of each other, who discovered a method using electrolysis. The two scientists were Charles Martin Hall (USA), who discovered the process in 1886, and the French chemist Paul Héroult, who discovered the process independently in the same year. The process they developed, often called the Hall–Héroult process, involves the electrolysis of aluminium oxide (alumina). The process involves the following stages.

- Bauxite, an impure form of aluminium oxide, is first treated with sodium hydroxide to obtain pure aluminium oxide, removing impurities such as iron(III) oxide and sand. This improves the conductivity of the molten aluminium oxide.
- The purified aluminium oxide is then dissolved in molten cryolite ( $\text{Na}_3\text{AlF}_6$ ). Cryolite, a mineral found naturally in Greenland, is used to reduce the working temperature of the Hall–Héroult cell from  $2017^\circ\text{C}$  (the melting point of pure aluminium oxide) to between  $800$  and  $1000^\circ\text{C}$ . Therefore, the cryolite provides a considerable



Figure 5.5 Bauxite mining.



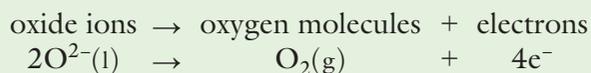
**Figure 5.6** The Hall-Héroult cell is used in industry to extract aluminium by electrolysis.

saving in the energy requirements of the process. In recent years it has become necessary to manufacture the cryolite.

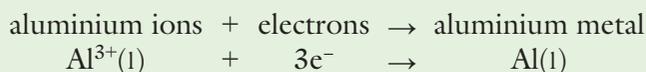
- The molten mixture is then electrolysed in a cell similar to that shown in Figure 5.6.

The anodes of this process are blocks of graphite which are lowered into the molten mixture from above. The cathode is the graphite lining of the steel vessel containing the cell.

Aluminium oxide is an ionic compound. When it is melted the ions become mobile, as the strong electrostatic forces of attraction between them are broken by the input of heat energy. During electrolysis the negatively charged oxide ions are attracted to the anode (the positive electrode), where they lose electrons (oxidation).

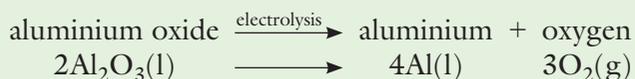


The positive aluminium ions are attracted to the cathode (the negative electrode). They gain electrons to form molten aluminium metal (reduction).

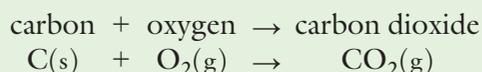


A handy way of remembering it is **OIL RIG** (**O**xidation **I**s **L**oss, **R**eduction **I**s **G**ain of electrons).

The overall reaction which takes place in the cell is:



The molten aluminium collects at the bottom of the cell and it is siphoned out at regular intervals. No problems arise with other metals being deposited, since the cryolite is largely 'unaffected' by the flow of electricity. Problems do arise, however, with the graphite anodes. At the working temperature of the cell, the oxygen liberated reacts with the graphite anodes, producing carbon dioxide.



The anodes burn away and have to be replaced on a regular basis.

The electrolysis of aluminium oxide is a continuous process in which vast amounts of electricity are used. Approximately 15 kWh of electricity are used to produce 1 kg of aluminium. In order to make the process an economic one, a cheap form of electricity is required. Hydroelectric power (HEP) is usually used for this process. The plant shown in Figure 5.7 uses an HEP scheme to provide some of the electrical energy required for this process. Further details about HEP are given in Chapter 6, p. 94.



**Figure 5.7** An aluminium smelting plant.

Using cheap electrical energy has allowed aluminium to be produced in such large quantities that it is the second most widely used metal after iron. It is used in the manufacture of electrical cables because of its low density, chemical inertness and good electrical conductivity. Owing to the first two of these properties, it is also used for making cars, bikes, cooking foil and food containers as well as in alloys (Chapter 10, p. 167) such as duralumin, which is used in the manufacture of aeroplane bodies (Figure 5.8). Worldwide production of aluminium now exceeds 40 million tonnes each year.



Figure 5.8 Aluminium is used in the manufacture of aeroplane bodies.

Environmental problems associated with the location of aluminium plants are concerned with:

- the effects of the extracted impurities, which form a red mud (Figure 5.9)
- the fine cryolite dust, which is emitted through very tall chimneys so as not to affect the surrounding area



Figure 5.9 Bauxite pollution of the Amazon in Brazil.

- the claimed link between environmental aluminium and a degenerative brain disorder called Alzheimer's disease – it is thought that aluminium is a major influence on the early onset of this disease. However, the evidence is still inconclusive.

### Questions

- 1 Produce a flow chart to summarise the processes involved in the extraction of aluminium metal from bauxite.
- 2 Explain why the mixture of gases formed at the anode contains oxygen, carbon dioxide and traces of fluorine.

### Anodising

This is a process in which the surface coating of oxide on aluminium ( $\text{Al}_2\text{O}_3$ ) is made thicker. In this process the aluminium object is made the anode in a cell in which the electrolyte is dilute sulfuric acid. During the electrolysis process, oxygen is produced at the anode and combines with the aluminium. The oxide layer on the surface of the aluminium therefore increases. Dyes can be mixed with the electrolyte and so the new thicker coating of oxide is colourful and also decorative (Figure 5.10).



**Figure 5.10** The oxide layer on the surface of these aluminium cups has been thickened, and dyes added to obtain the vibrant colours.

### Question

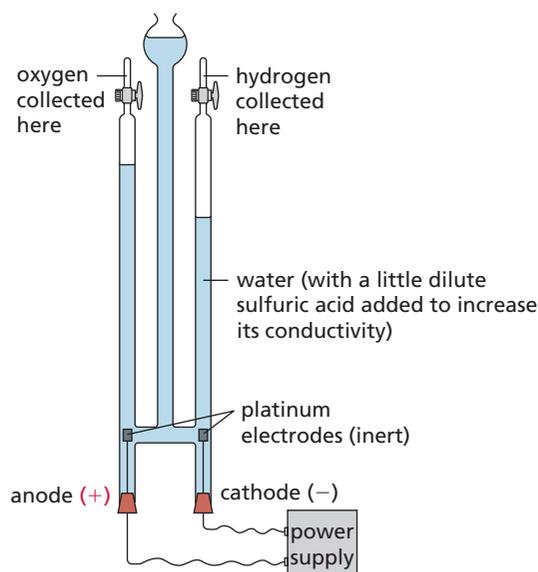
- 1 A student carries out the electrolysis of molten lead(II) iodide in a fume cupboard.
  - a Draw a diagram to show a suitable apparatus the student could use to carry out this experiment.
  - b Write anode and cathode reactions to represent the processes taking place during the electrolysis.
  - c Why does this experiment need to be carried out in a fume cupboard?
  - d Find uses for the anode product of this cell.

## ● Electrolysis of aqueous solutions

Other industrial processes involve the electrolysis of aqueous solutions. To help you to understand what is happening in these processes, we will first consider the electrolysis of dilute sulfuric acid.

### Electrolysis of dilute sulfuric acid

Pure water is a very poor conductor of electricity because there are so few ions in it. However, it can be made to decompose if an electric current is passed through it in a Hofmann voltameter, as in Figure 5.11.

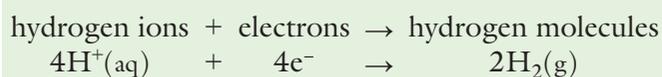


**Figure 5.11** A Hofmann voltameter used to electrolyse water.

To enable water to conduct electricity better, some dilute sulfuric acid (or sodium hydroxide solution) is added. When the power is turned on and an electric current flows through this solution, gases can be seen to be produced at the two electrodes and they are collected in the side arms of the apparatus. After about 20 minutes, roughly twice as much gas is produced at the cathode as at the anode.

The gas collected at the cathode burns with a squeaky pop, showing it to be hydrogen gas (Table 16.6, p. 263).

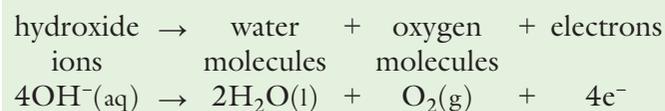
For hydrogen to be collected in this way, the positively charged hydrogen ions must have moved to the cathode.



If during this process the water molecules lose  $\text{H}^+(\text{aq})$ , then the remaining portion must be hydroxide ions,  $\text{OH}^-(\text{aq})$ . These ions are attracted to the anode.

The gas collected at the anode relights a glowing splint, showing it to be oxygen (Table 16.6, p. 263).

This gas is produced in the following way.



This experiment was first carried out by Sir Humphry Davy. It confirmed that the formula for water was  $\text{H}_2\text{O}$ .

It should be noted that in the electrolysis of dilute sulfuric acid, platinum (an inert electrode) may be replaced by carbon (graphite). The only difference to occur is that as well as oxygen being produced at the anode, a little carbon dioxide will also be formed.

## The chlor-alkali industry

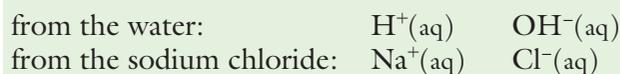
The electrolysis of saturated sodium chloride solution (brine) is the basis of a major industry. In countries where rock salt (sodium chloride) is found underground it is mined. In other countries it can be obtained by evaporation of sea water in large shallow lakes. Three very important substances are produced

in this electrolysis process – chlorine, sodium hydroxide and hydrogen. The electrolytic process is a very expensive one, requiring vast amounts of electricity. The process is economical only because all three products have a large number of uses (Figure 5.12).

There are two well-established methods for electrolysis of brine, the **diaphragm cell** and the **mercury cell**. However, recent developments in electrolysis technology, by chemical engineers, have produced the **membrane cell** (Figure 5.13). This method is now preferred to the other two because it produces a purer product, it causes less pollution and it is cheaper to run.

The brine is first purified to remove calcium, strontium and magnesium compounds by a process of ion exchange (see Chapter 11, p. 189).

The membrane cell is used continuously, with fresh brine flowing into the cell as the process breaks up the brine. The cell has been designed to ensure that the products do not mix. The ions in this concentrated sodium chloride solution are:



When the current flows, the chloride ions,  $\text{Cl}^-(\text{aq})$ , are attracted to the anode. Chlorine gas is produced by the electrode process.



This leaves a high concentration of sodium ions,  $\text{Na}^+(\text{aq})$ , around the anode.

The hydrogen ions,  $\text{H}^+(\text{aq})$ , are attracted to the cathode and hydrogen gas is produced.



This leaves a high concentration of hydroxide ions,  $\text{OH}^-(\text{aq})$ , around the cathode. The sodium ions,  $\text{Na}^+(\text{aq})$ , are drawn through the membrane, where they combine with the  $\text{OH}^-(\text{aq})$  to form sodium hydroxide,  $\text{NaOH}$ , solution. The annual production worldwide is now in excess of 60 million tonnes.

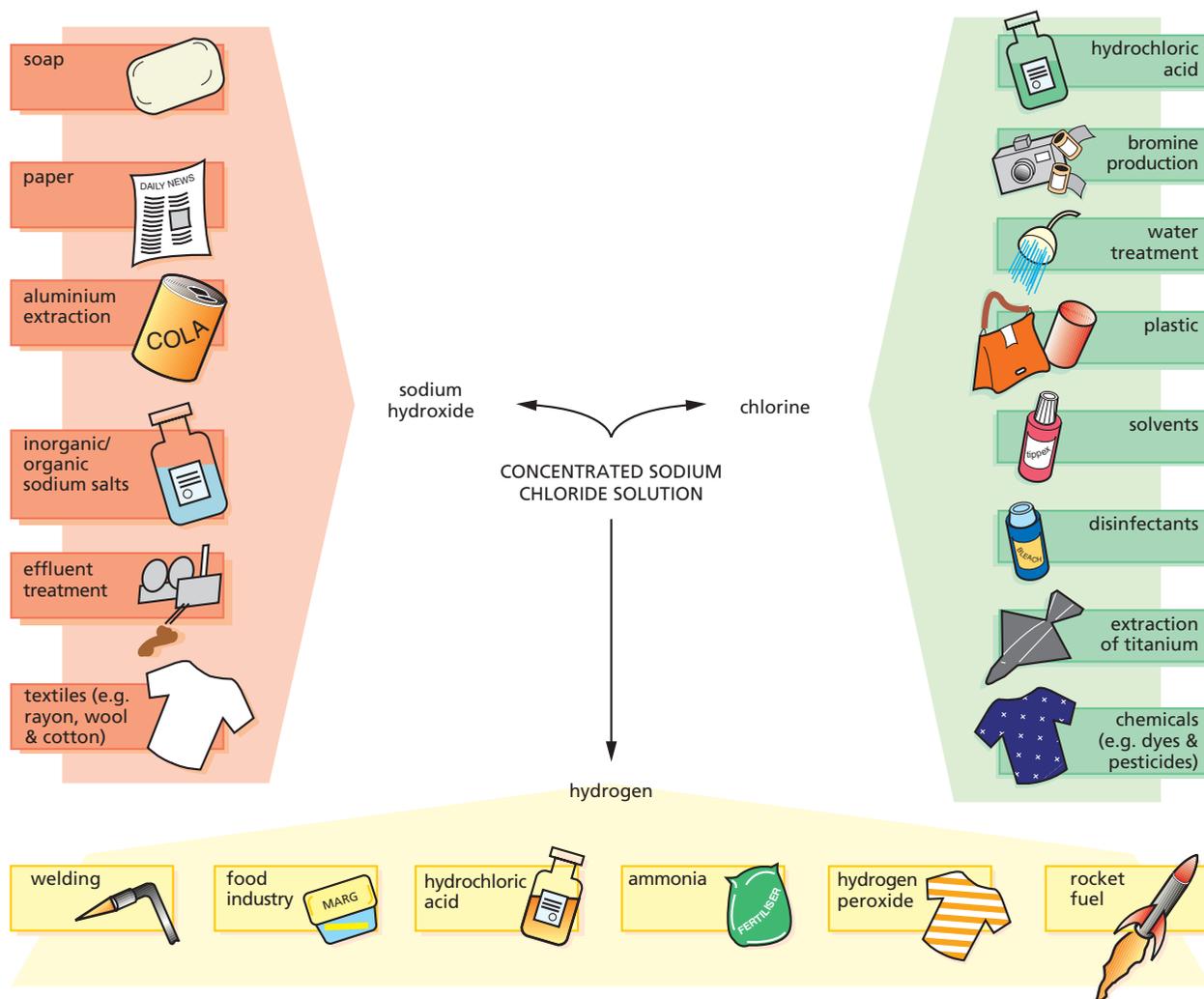
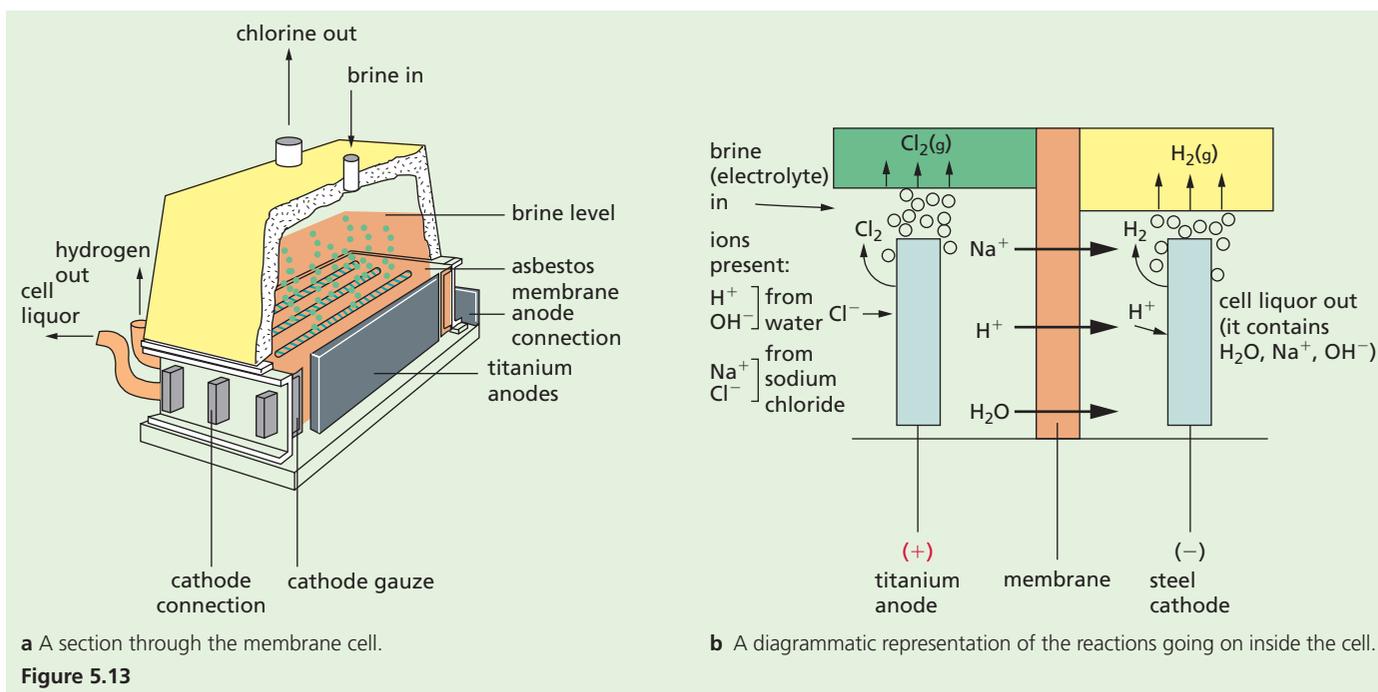


Figure 5.12 The chlor-alkali industry.



a A section through the membrane cell.

Figure 5.13

b A diagrammatic representation of the reactions going on inside the cell.

## Tests for aqueous cations

Aqueous sodium hydroxide can be used to identify salts of  $\text{Al}^{3+}$ ,  $\text{Ca}^{2+}$ ,  $\text{Cr}^{3+}$ ,  $\text{Cu}^{2+}$ ,  $\text{Fe}^{2+}$ ,  $\text{Fe}^{3+}$  and  $\text{Zn}^{2+}$ . The colour of the precipitate and its behaviour in excess sodium hydroxide solution will help identify the metal present. In the case of ammonium salts, ammonia gas is produced on warming.

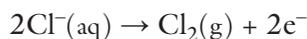
Table 5.1

Cation solution	Effect of aqueous sodium hydroxide
Aluminium ( $\text{Al}^{3+}$ )	White precipitate, soluble in excess giving a colourless solution
Ammonium ( $\text{NH}_4^+$ )	Ammonia produced on warming
Calcium ( $\text{Ca}^{2+}$ )	White precipitate, insoluble in excess
Chromium ( $\text{Cr}^{3+}$ )	Green precipitate, soluble in excess
Copper ( $\text{Cu}^{2+}$ )	Light blue precipitate, insoluble in excess
Iron(II) ( $\text{Fe}^{2+}$ )	Green precipitate, insoluble in excess
Iron(III) ( $\text{Fe}^{3+}$ )	Red-brown precipitate, insoluble in excess
Zinc ( $\text{Zn}^{2+}$ )	White precipitate, soluble in excess, giving a colourless solution

If the nitrate ion is present in solution, then upon addition of aqueous sodium hydroxide and a little aluminium foil, ammonia gas is produced on warming.

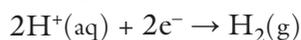
## Electrolysis of concentrated hydrochloric acid

When concentrated hydrochloric acid is electrolysed, two gases are produced. Both chloride ions and hydroxide ions would be attracted to the anode. The chloride ions are at a much higher concentration than the hydroxide ions and go on to produce chlorine gas by losing electrons (oxidation).



The chlorine gas can be tested by placing a piece of damp blue litmus paper near the anode. The paper will first turn red and then it will be bleached (turned white).

The only positively charged ion is the hydrogen ion. These are attracted to the cathode where they produce hydrogen gas by the gain of electrons (reduction).



Hydrogen gas, if collected from the cathode, will give a 'squeaky pop' when mixed with air and ignited.

## Questions

- Suggest a reason for only 'roughly' twice as much hydrogen gas being produced at the cathode as oxygen gas at the anode in the electrolysis of water.
- Account for the following observations which were made when concentrated sodium chloride solution, to which a little universal indicator had been added, was electrolysed in the laboratory in a Hofmann voltameter.
  - The universal indicator initially turns red in the region of the anode, but as the electrolysis proceeds it loses its colour.
  - The universal indicator turns blue in the region of the cathode.
- Why is it important to remove compounds of calcium, strontium and magnesium before brine is electrolysed?
- The uses of sodium hydroxide can be separated on a percentage basis as follows:
 

Neutralisation	5%
Paper manufacture	5%
Oil refining	5%
Soap/detergents	5%
Manufacture of rayon and acetate fibres	16%
Manufacture of chemicals	30%
Miscellaneous uses	34%

 Use a graph-plotting program to create a 'pie' chart of these data.

## Electrolysis of copper(II) sulfate solution

Copper(II) sulfate solution ( $\text{CuSO}_4(\text{aq})$ ) may be electrolysed using inert graphite electrodes in a cell similar to that shown in Figure 5.14. When the solution is electrolysed, oxygen gas and copper metal are formed at the anode and cathode respectively. Four ions are present in solution:

from the water:  $\text{H}^+(\text{aq})$      $\text{OH}^-(\text{aq})$   
 from the copper(II) sulfate:  $\text{Cu}^{2+}(\text{aq})$      $\text{SO}_4^{2-}(\text{aq})$

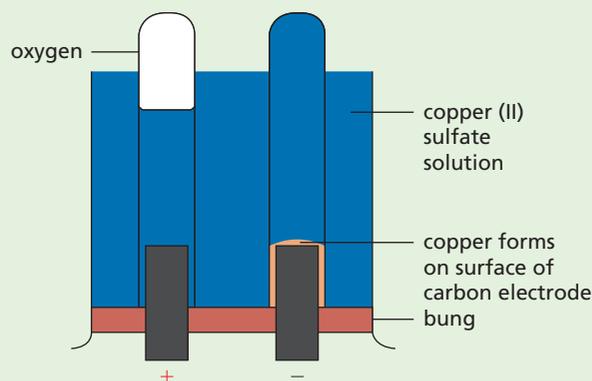
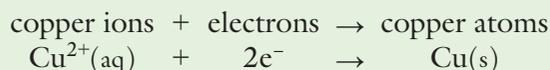
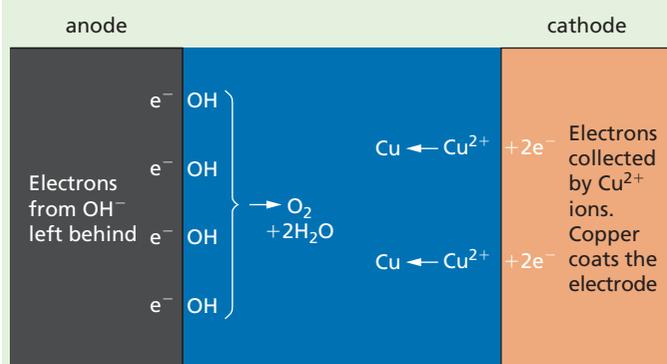
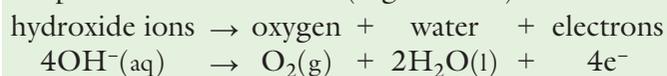


Figure 5.14 The electrolysis of copper(II) sulfate solution using inert electrodes.

$\text{H}^+(\text{aq})$  and  $\text{Cu}^{2+}(\text{aq})$  ions are both attracted to the cathode, the  $\text{Cu}^{2+}$  ions accepting electrons more readily than the  $\text{H}^+$  ions (preferential discharge). Copper metal is therefore deposited at the cathode (Figure 5.14).



$\text{OH}^-(\text{aq})$  and  $\text{SO}_4^{2-}(\text{aq})$  ions are both attracted to the anode. The  $\text{OH}^-$  ions release electrons more easily than the  $\text{SO}_4^{2-}$  ions, so oxygen gas and water are produced at the anode (Figure 5.15).

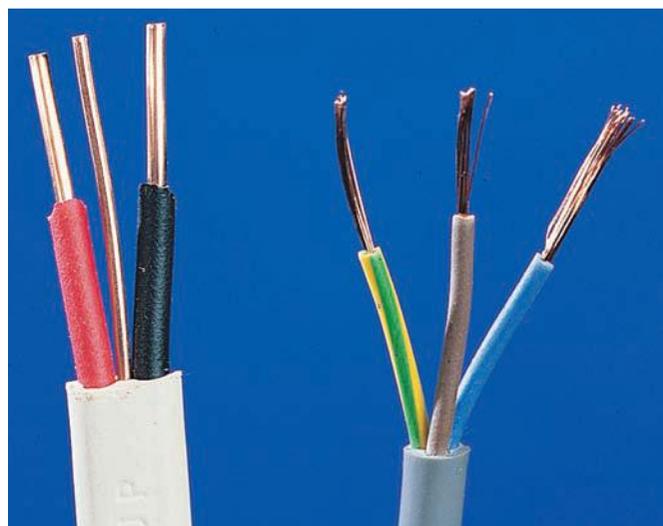


**Figure 5.15** Oxygen is given off at the anode and copper is deposited at the cathode.

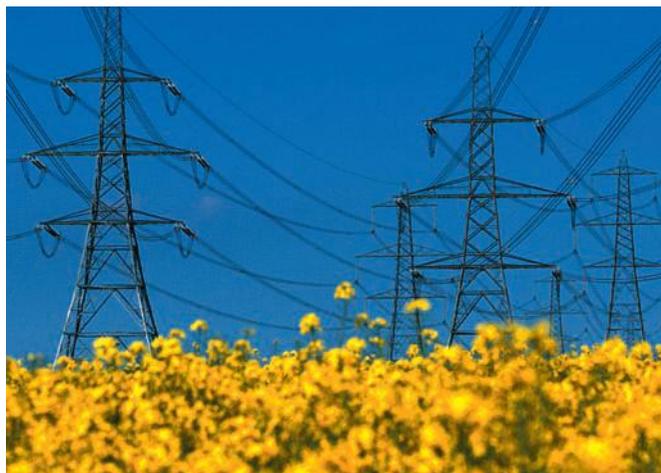
## Purification of copper

Because copper is a very good conductor of electricity, it is used for electrical wiring and cables (Figure 5.16). Pure copper is also used in the manufacture of cooking utensils owing to its high thermal conductivity, a property of its metallic structure (Chapter 3, p. 55).

However, even small amounts of impurities cut down this conductivity quite noticeably whether in fine wires or larger cables. The metal must be 99.99% pure to be used in this way. To ensure this level of purity, the newly extracted copper has to be purified by electrolysis.

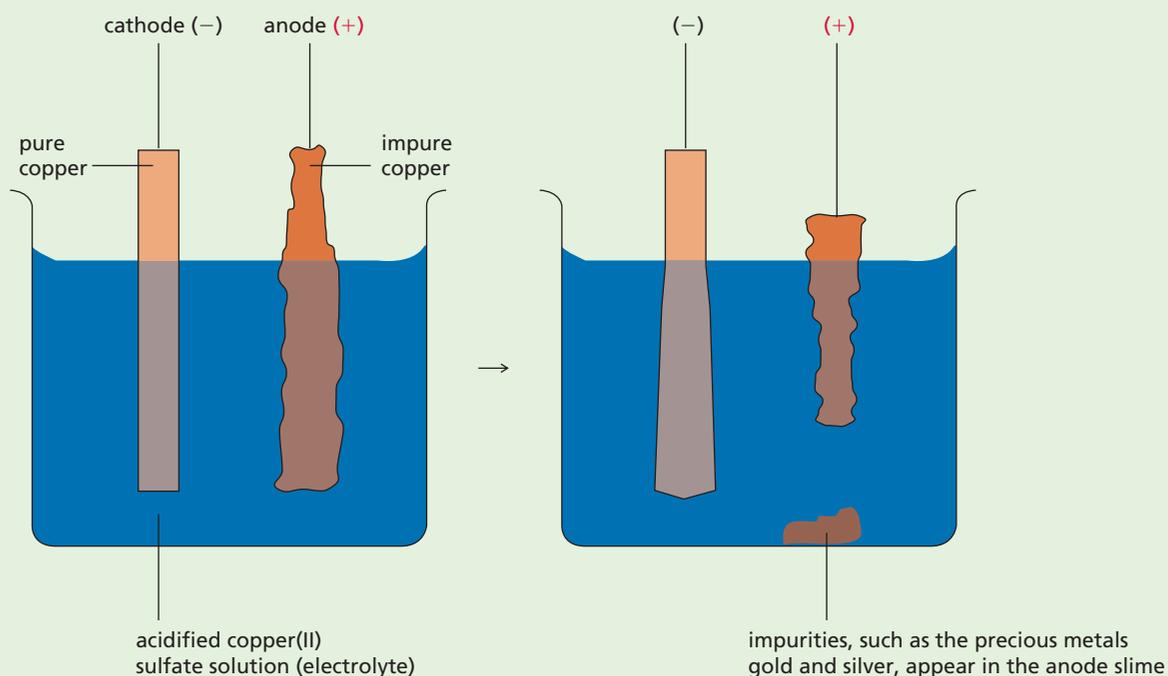


**Figure 5.16a** The copper used in electrical wiring has to be very pure.



**Figure 5.16b** Due to the high density of copper and its cost, steel-cored aluminium cables are used for electrical transmission through national grids.

The impure copper is used as the anode and is typically 1 m square, 35–50 mm thick and 330 kg in weight. The cathode is a 1 mm thick sheet and weighs about 5 kg; it is made from very pure copper. Because copper is itself involved in the electrolytic process, the copper cathode is known as an ‘active’ electrode. The electrolyte is a solution of copper(II) sulfate ( $0.3 \text{ mol dm}^{-3}$ ) acidified with a  $2 \text{ mol dm}^{-3}$  solution of sulfuric acid to help the solution conduct electricity (Figure 5.17).



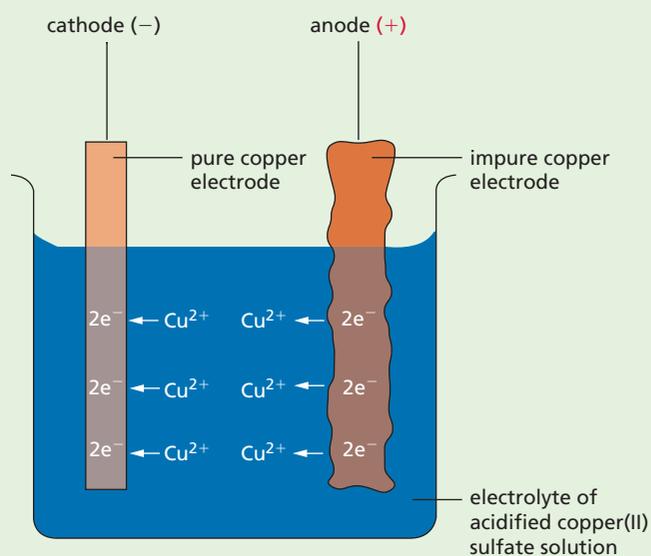
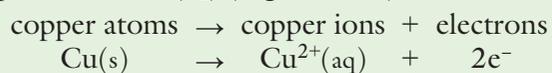
**Figure 5.17** Copper purification process.

When the current flows, the copper moves from the impure anode to the pure cathode. Any impurities fall to the bottom of the cell and collect below the anode in the form of a slime. This slime is rich in precious metals and the recovery of these metals is an important aspect of the economics of the process. The electrolysis proceeds for about three weeks until the anodes are reduced to about 10% of their original size and the cathodes weigh between 100 and 120 kg. A potential of 0.25 V and a current density of  $200 \text{ A m}^{-2}$  are usually used.

The ions present in the solution are:

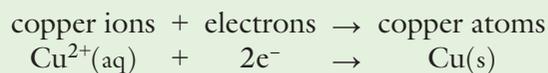
from the water:  $\text{H}^+(\text{aq})$        $\text{OH}^-(\text{aq})$   
 from the copper(II) sulfate:  $\text{Cu}^{2+}(\text{aq})$        $\text{SO}_4^{2-}(\text{aq})$

During the process the impure anode loses mass because the copper atoms lose electrons and become copper ions,  $\text{Cu}^{2+}(\text{aq})$  (Figure 5.18).



**Figure 5.18** The movement of ions in the purification of copper by electrolysis.

The electrons released at the anode travel around the external circuit to the cathode. There the electrons are passed on to the copper ions,  $\text{Cu}^{2+}(\text{aq})$ , from the copper(II) sulfate solution and the copper is deposited or copper plated on to the cathode.



The annual production of copper worldwide is in excess of 16 million tonnes. However, a large amount of the copper we need is obtained by recycling. This way of obtaining copper is increasing in importance as it becomes more difficult and expensive to locate and extract the copper ore.

## ● Electrolysis guidelines

The following points may help you work out the products of electrolysis in unfamiliar situations. They will also help you remember what happens at each electrode.

- Non-metals are produced at the anode whereas metals and hydrogen gas are produced at the cathode.
- At the inert anode, chlorine, bromine and iodine (the halogens) are produced in preference to oxygen.
- At the inert cathode, hydrogen is produced in preference to metals unless unreactive metals such as copper and nickel are present.

### Questions

- 1 Why do you think it is advantageous to use inert electrodes in the electrolysis processes?
- 2 Predict the products of electrolysis of a solution of copper(II) sulfate if carbon electrodes are used instead of those made from copper as referred to in the purification of copper section.
- 3 Predict the products of the electrolysis of concentrated hydrochloric acid using platinum electrodes.
- 4 Using your knowledge of electrolysis, predict the likely products of the electrolysis of copper(II) chloride solution, using platinum electrodes. Write electrode equations for the formation of these products.
- 5 Suggest reasons why:
  - a copper and aluminium are used in electrical cables
  - b plastics and ceramics are used to manufacture handles for metallic cooking utensils.

## ● Electroplating

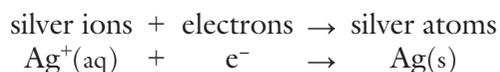
Electroplating is the process involving electrolysis to plate, or coat, one metal with another or a plastic with a metal. Often the purpose of electroplating is to give a protective coating to the metal beneath. For example, bath taps are chromium plated to prevent corrosion, and at the same time are given a shiny, more attractive finish (Figure 5.19).

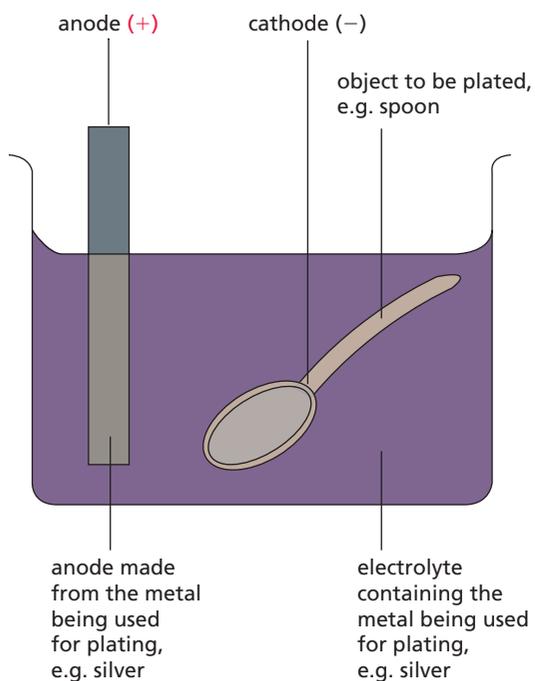


**Figure 5.19** This tap has been chromium plated.

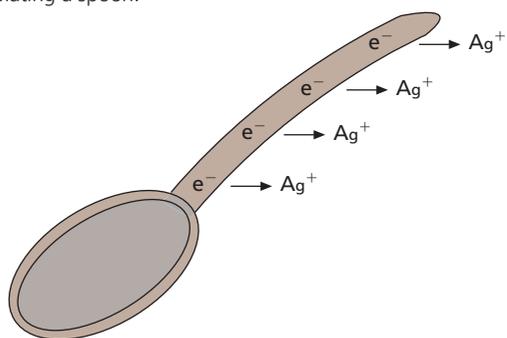
The electroplating process is carried out in a cell such as the one shown in Figure 5.20a (p. 84). This is often known as the ‘plating bath’ and it contains a suitable electrolyte, usually a solution of a metal salt.

For silver plating the electrolyte is a solution of a silver salt. The article to be plated is made the cathode in the cell so that the metal ions move to it when the current is switched on. The cathode reaction in this process is:

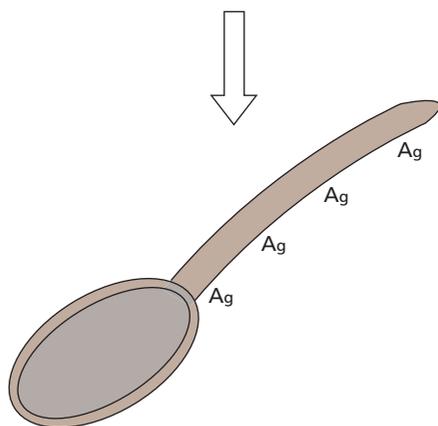




a Silver plating a spoon.



The  $\text{Ag}^+$  ions are attracted to the cathode, where they gain electrons



A coating of silver forms on the spoon at the cathode

b Explaining silver plating.

Figure 5.20

### Plating plastics

Nowadays it is not only metals that are electroplated. Plastics have been developed that are able to conduct electricity. For example, the plastic poly(pyrrole) can be electroplated in the same way as the metals we have discussed earlier (Figure 5.21).



Figure 5.21 This plastic has been coated with copper by electrolysis.



Figure 5.22 This leaf has been electroplated.

## Questions

- 1 The leaf in Figure 5.22 has been copper plated. Suggest a suitable method for copper plating the leaf.
- 2 Explain why copper(II) chloride solution would not be used as an electrolyte in the electrolyte cell used for copper plating.
- 3 Write equations which represent the discharge at the cathode of the following ions:
  - a  $K^+$
  - b  $Pb^{2+}$
  - c  $Al^{3+}$and at the anode of:
  - d  $Br^-$
  - e  $O^{2-}$
  - f  $F^-$

## Checklist

After studying Chapter 5 you should know and understand the following terms.

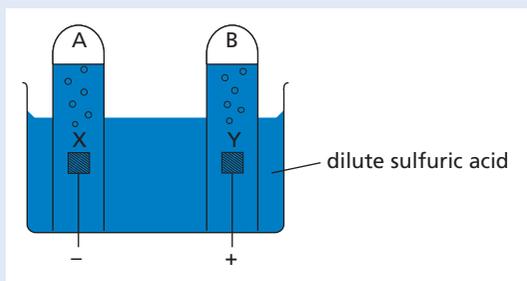
- **Anions** Negative ions; these are attracted to the anode.
- **Anode** The positive electrode. It is positively charged because electrons are drawn away from it.
- **Cathode** The negative electrode. It is negatively charged because an excess of electrons move towards it.
- **Cations** Positive ions; these are attracted to the cathode.
- **Electrode** A point where the electric current enters and leaves the electrolytic cell. An inert electrode is usually made of platinum or carbon and does not react with the electrolyte or the substances produced at the electrodes themselves.
- **Electrolysis** A process in which a chemical reaction is caused by the passage of an electric current.
- **Electrolyte** A substance which will carry electric current only when it is molten or dissolved.
- **Electroplating** The process of depositing metals from solution in the form of a layer on other surfaces such as metal or plastic.
- **Hall-Héroult cell** The electrolysis cell in which aluminium is extracted from purified bauxite dissolved in molten cryolite at  $900^\circ\text{C}$ . This cell has both a graphite anode and a graphite cathode.
- **Inert electrode** These are electrodes that do not react with the products of electrolysis, e.g. carbon, platinum.
- **Membrane cell** An electrolytic cell used for the production of sodium hydroxide, hydrogen and chlorine from brine in which the anode and cathode are separated by a membrane.
- **Oxidation** Takes place at the anode and involves a negative ion losing electrons.
- **Reduction** Takes place at the cathode and involves a positive ion gaining electrons.



# Electricity and chemistry

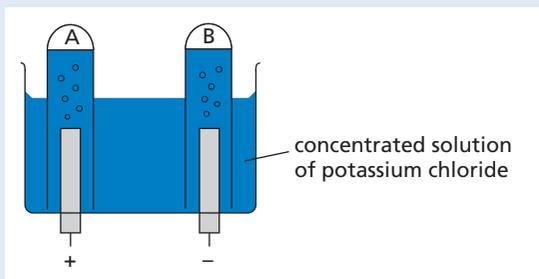
## Additional questions

1



This is a diagram of an experiment in which electricity was passed through a mixture of distilled water containing a little dilute sulfuric acid.

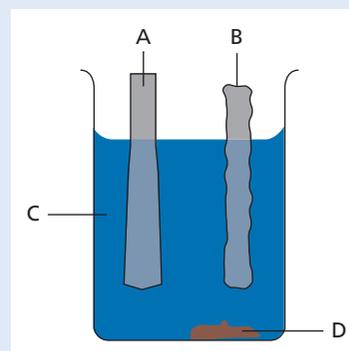
- Name the gas that collects at **A**.
  - Name the gas that collects at **B**.
  - If  $100\text{ cm}^3$  of gas collects in **A** how much would there be in **B**?
  - Name the metal usually used for **X** and **Y**.
  - X** is called the \_\_\_\_\_.
  - Y** is called the \_\_\_\_\_.
  - Write down the formulae of the three ions present in the solution.
  - Write down the equations for the reactions that take place at both **X** and **Y** (or describe the changes that take place if you cannot write the equations).
- 2 The apparatus shown in the diagram below was used to investigate the gases produced when a concentrated solution of potassium chloride was electrolysed.



- Name a non-metal suitable for use as electrodes.
- Name the gas collected in **A** and the gas collected in **B**.
- Describe how you would test the gases collected.
- The volume of gas collected in **B** was slightly less than that collected in **A**. The teacher said the volumes should have been equal but gave

a simple explanation of the 'missing' gas in **B**. What was the explanation? (Assume that the apparatus was working perfectly).

- Write down the equations which describe the production of the gases at the electrodes in **A** and **B**.
  - (i) If the concentrated solution of potassium chloride was now replaced by dilute sodium hydroxide what gases would be produced at **A** and **B**?  
(ii) In what ratio would you expect these gases to be produced?
- 3 Explain the meaning of each of the following terms. Use a suitable example, in each case, to help with your explanation.
- |                 |              |
|-----------------|--------------|
| a Anode.        | e Anion.     |
| b Cathode.      | f Cation.    |
| c Electrolysis. | g Oxidation. |
| d Electrolyte.  | h Reduction. |
- 4 Copper is purified by electrolysis, as in the example shown below.



- Name the materials used for the electrodes **A** and **B**.
- Name the electrolyte **C** and substance **D**.
- Why is substance **D** of economic importance in respect of this process?
- Write equations for the reactions which take place at the cathode and anode during this process.
- Draw a labelled diagram to show the cell after electrolysis has taken place.
- Why has electrolyte **C** to be acidified with the dilute sulfuric acid?
- Why does copper have to be 99.99% pure for use in electrical cables?

- 5 Copy and complete the table below, which shows the results of the electrolysis of four substances using inert electrodes.

Electrolyte	Product at anode (positive electrode)	Product at cathode (negative electrode)
Molten aluminium oxide		Aluminium
Concentrated sodium chloride solution	Chlorine	
Molten lithium chloride		
Silver nitrate solution		Silver

- State what you understand by 'inert electrodes'.
  - Explain why the lithium chloride solution becomes progressively more alkaline during electrolysis.
  - Explain why solid lithium chloride is a non-conductor of electricity, whereas molten lithium chloride and lithium chloride solution are good conductors of electricity.
  - During the electrolysis of molten aluminium chloride ( $\text{AlCl}_3$ ) the carbon anodes are burned away. Explain why this should happen and write balanced chemical equations for the reactions that take place.
- 6 A pupil carried out an experiment in a fume cupboard to find out how electricity affected different substances. Some of the substances were in aqueous solution, others were in the molten state. Carbon electrodes were used in each experiment and she wrote down her results in a table with these headings.

Substance	What was formed at the cathode (-)	What was formed at the anode (+)

Make a table like the one shown and fill it in with what you think happened for each of the substances below.

- Molten lead iodide.
- Sugar solution.
- Silver nitrate solution.
- Copper(II) sulfate solution.
- Molten sodium bromide.
- Ethanol solution.

- 7 Sodium hydroxide is made by the electrolysis of brine.
- Draw and label a simplified diagram of the cell used in this process. Make certain that you have labelled on the diagram:
    - the electrolyte
    - the material of the electrodes
    - the material of the membrane.
  - Write equations for the reactions which take place at the cathode and anode. State clearly whether a reaction is oxidation or reduction.
  - Give two large-scale uses of the products of this electrolytic process.
  - Comment on the following statement: 'This electrolytic process is a very expensive one'.
  - Both the membrane cell and the older mercury cell make sodium hydroxide of high purity. Explain why the membrane cell is now the preferred way of making sodium hydroxide.
- 8 Electroplating is an important industrial process.
- Explain what electroplating is.
  - Why are certain metals electroplated?
  - Give two examples of the use of electroplating.

## 6

# Chemical energetics

## Substances from oil

Oil refining

## Fossil fuels

### What is a fuel?

How are fossil fuels used?

## Alternative sources of energy

Nuclear power

Hydroelectric power

Biomass and biogas

Hydrogen – the fuel of the future

## Chemical energy

Combustion

Enthalpy of neutralisation (molar heat of neutralisation)

## Changes of state

## Cells and batteries

Fuel cells

## Checklist

## Additional questions

## Substances from oil

What do the modes of transport shown in Figure 6.1 have in common? They all use liquids obtained from **crude oil** as fuels.



**Figure 6.1** These modes of transport all use fuels that have been extracted from crude oil.

## Oil refining

Crude oil is a complex mixture of compounds known as **hydrocarbons** (Figure 6.2). Hydrocarbons are molecules which contain only the elements carbon and hydrogen bonded together covalently (Chapter 3, p. 46). These carbon compounds form the basis of a group called **organic compounds**. All living things are made from organic compounds based on chains of carbon atoms similar to those found in crude oil. Crude oil is not only a major source of fuel but is also a raw material of enormous importance. It supplies a large and diverse chemical industry to make dozens of products.



Figure 6.2 Crude oil is a mixture of hydrocarbons.



Figure 6.3 An oil refinery.

Crude oil is not very useful to us until it has been processed. The process, known as **refining**, is carried out at an oil refinery (Figure 6.3).

Refining involves separating crude oil into various batches or **fractions**. Chemists use a technique called **fractional distillation** to separate the different fractions. This process works in a similar way to that discussed in Chapter 2, p. 21, for separating ethanol (alcohol) and water. The different components (fractions) separate because they have different boiling points. The crude oil is heated to about  $400^{\circ}\text{C}$  to vaporise all the different parts of the mixture. The mixture of vapours passes into the fractionating column near the bottom (Figure 6.4, p. 90). Each fraction is obtained by collecting hydrocarbon molecules which have a boiling point in a given range of temperatures. For example, the fraction we know as petrol contains molecules which have boiling points between  $30^{\circ}\text{C}$  and  $110^{\circ}\text{C}$ . The molecules in this fraction contain between five and ten carbon atoms. These smaller molecules with lower boiling points condense higher up the tower. The bigger hydrocarbon molecules which have the higher boiling points condense in the lower half of the tower.

The liquids condensing at different levels are collected on **trays**. In this way the crude oil is separated into different fractions. These fractions usually contain a number of different hydrocarbons. The individual single hydrocarbons can then be obtained, again by refining the fraction by further distillation.

It is important to realise that the uses of the fractions depend on their properties. For example, one of the lower fractions, which boils in the range  $250\text{--}350^{\circ}\text{C}$ , is quite thick and sticky and makes a good lubricant. However, the petrol fraction burns very easily and this therefore makes it a good fuel for use in engines.

### Questions

- 1 What do you understand by the term hydrocarbon?
- 2 All organisms are composed of compounds which contain carbon. Why do you think carbon chemistry is often called 'organic chemistry'?
- 3 List the main fractions obtained by separating the crude oil mixture and explain how they are obtained in a refinery.

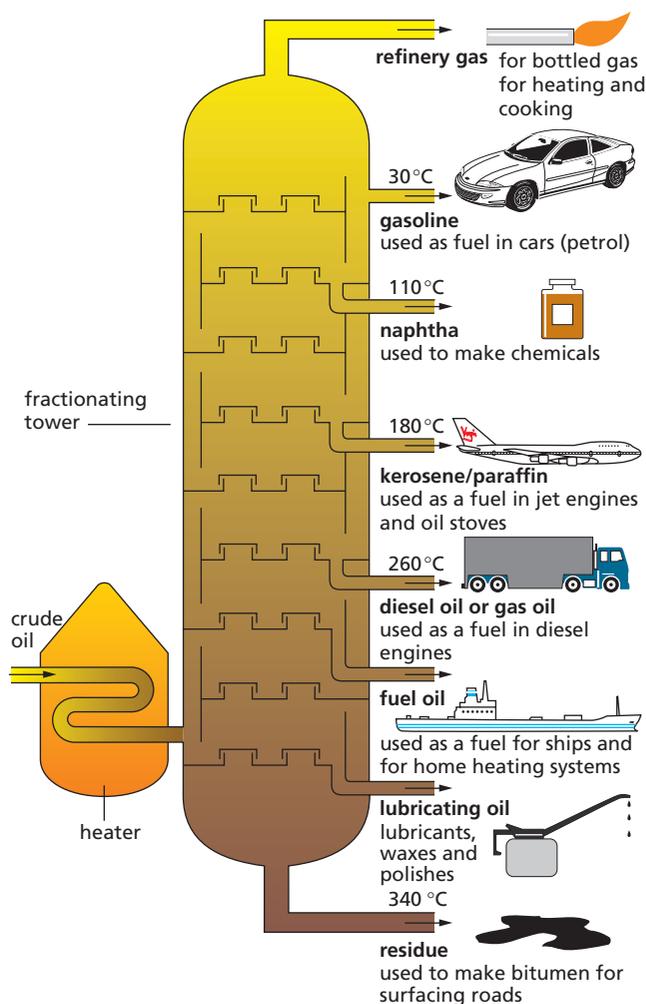


Figure 6.4a Fractional distillation of crude oil in a refinery.

## Fossil fuels

Coal, oil and natural gas are all examples of **fossil fuels**. The term fossil fuels is derived from the fact that they are formed from dead plants and animals which were fossilised over 200 million years ago during the carboniferous era.

Coal was produced by the action of pressure and heat on dead wood from ancient forests which once grew in the swampland in many parts of the world under the prevailing weather conditions of that time. When dead trees fell into the swamps they were buried by mud. This prevented aerobic decay (which takes place in the presence of oxygen). Over millions of years, due to movement of the Earth's crust as well as to changes in climate, the land sank and the decaying wood became covered by even more layers of mud and sand. Anaerobic decay (which takes place in the absence of oxygen) occurred, and as time passed the gradually forming coal became more and more compressed as other material was laid down above it (Figure 6.5).



b Uses of the different fractions obtained from crude oil.



Figure 6.5 Piece of coal showing a fossilised leaf.

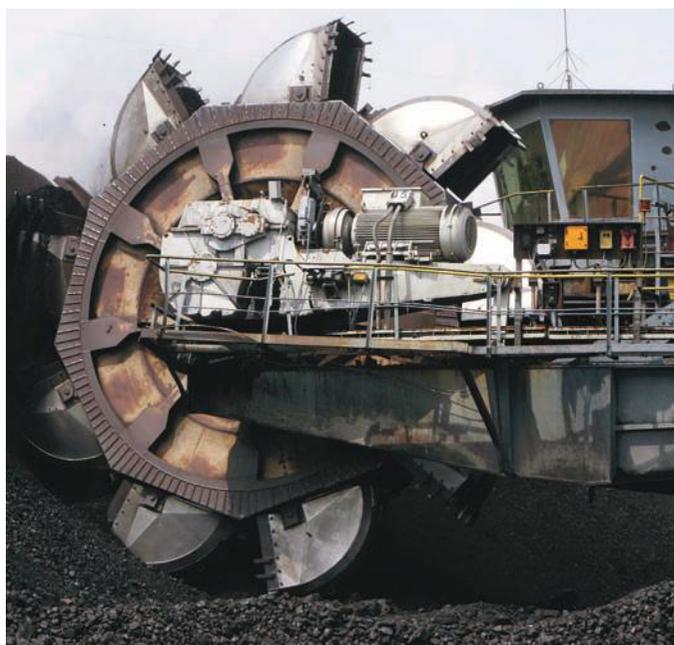


Figure 6.6 Cutting coal today is extremely mechanised.

Over millions of years, as the layers of forming coal were pushed deeper and the pressure and temperature increased, the final conversion to coal took place (Figure 6.6). Different types of coal were formed as a result of different pressures being applied during its formation. For example, anthracite is a hard coal with a high carbon content, typical of coal produced at greater depths. Table 6.1 shows some of the different types of coal along with their carbon contents.

Table 6.1 The different coal types.

Type of coal	Carbon content/%
Anthracite	90
Bituminous coal	60
Lignite	40
Peat	20

Oil and gas were formed during the same period as coal. It is believed that oil and gas were formed from the remains of plants, animals and bacteria that once lived in seas and lakes. This material sank to the bottom of these seas and lakes and became covered in mud, sand and silt which thickened with time.

Anaerobic decay took place and, as the mud layers built up, high temperatures and pressures were created which converted the material slowly into oil and gas. As rock formed, earth movements caused it to buckle and split, and the oil and gas were trapped in folds beneath layers of non-porous rock or cap-rock (Figures 6.7 and 6.8).



Figure 6.7 Oil production in the North Sea.

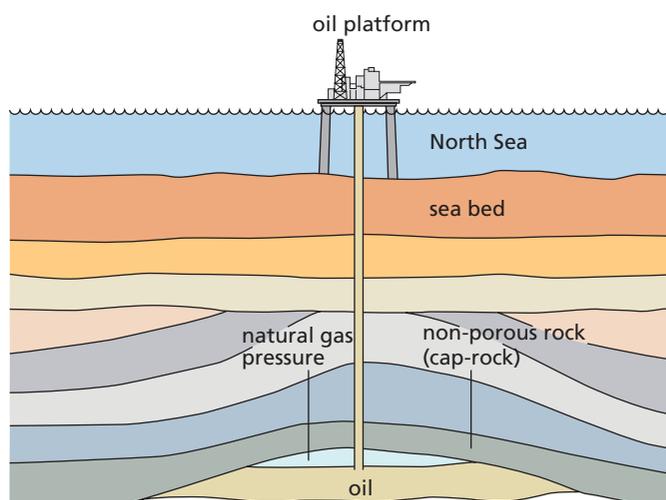


Figure 6.8 Natural gas and oil are trapped under non-porous rock.

### Questions

- Coal, oil and natural gas are all termed 'fossil fuels'. Why is the word 'fossil' used in this context?
- a Name the process by which plants convert carbon dioxide and water into glucose.  
b What conditions are necessary for this process to occur?
- Draw a flow diagram to represent the formation of coal, oil or gas.

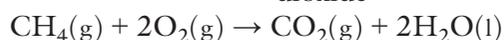
## ● What is a fuel?

A fuel is a substance which can be conveniently used as a source of energy. Fossil fuels release energy in the form of heat when they undergo **combustion**.

fossil fuel + oxygen → carbon dioxide + water + energy

For example, natural gas burns readily in air (Chapter 14, p. 220).

methane + oxygen → carbon + water + energy  
dioxide



It should be noted that natural gas, like crude oil, is a mixture of hydrocarbons such as methane (main constituent), ethane and propane, and may also contain some sulfur. The sulfur content varies from source to source. Natural gas obtained from the North Sea is quite low in sulfur.

The perfect fuel would be:

- cheap
- available in large quantities
- safe to store and transport
- easy to ignite and burn, causing no pollution
- capable of releasing large amounts of energy.

Solid fuels are safer than volatile liquid fuels like petrol and gaseous fuels like natural gas.

### Question

- 1 'We have not yet found the perfect fuel.' Discuss this statement.

### How are fossil fuels used?

A major use of fossil fuels is in the production of electricity. Coal, oil and natural gas are burned in power stations (Figure 6.9) to heat water to produce steam, which is then used to drive large turbines (Figure 6.10). A high proportion of the electricity generated worldwide is derived from fossil fuels. However, it should be noted that the relative importance of the three major fossil fuels is changing. Coal and oil are becoming less important while natural gas is increasingly important.

In a power station, the turbine drives a generator to produce electricity which is then fed into the national grid (Figure 6.10). The national grid is a system for distributing electricity throughout a country.



Figure 6.9 A Hongkong Electric Co. Power Station, Hong Kong, China.

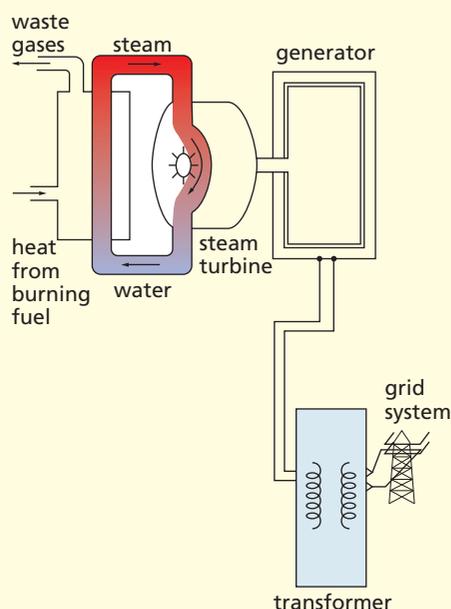


Figure 6.10 The way in which fuels are used to produce electricity.

Other major uses of the fossil fuels are:

- as major feedstock (raw material) for the chemicals and pharmaceuticals industries
- or domestic and industrial heating and cooking
- as fuels for various forms of vehicle transport.

### Question

- 1 'Fossil fuels are a major feedstock for the chemical and pharmaceutical industries.' With reference to Chapters 14 and 15, give examples which support this statement.

## Alternative sources of energy

Fossil fuels are an example of **non-renewable** resources, so called because they are not being replaced at the same rate as they are being used up. For example, we have approximately 60 years' supply of crude oil remaining, depending on the way in which it is extracted and the rate at which it is used, maybe more (Table 6.2). It is important to use non-renewable fuels very carefully and to consider alternative, **renewable** sources of energy for use now and in the future.

**Table 6.2** Estimates of how long our fossil fuels will last.

Fossil fuel	Estimated date it is expected to run out
Gas	2110
Oil	2080
Coal	2500

### Nuclear power

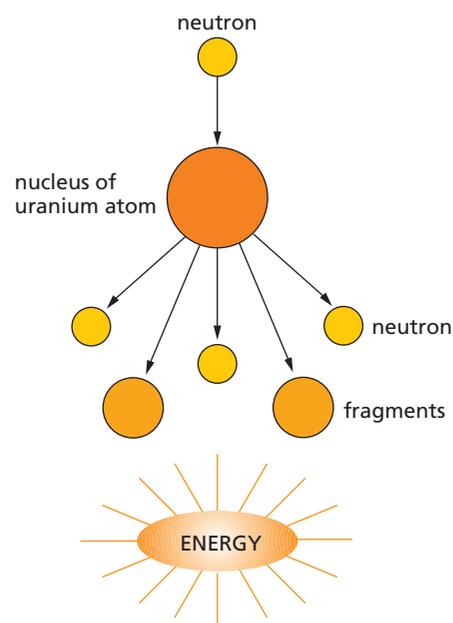
Calder Hall power station in Cumbria, UK, on the site of the present-day nuclear power complex at Sellafield (Figure 6.11), opened in 1956 and was the first nuclear reactor in the world to produce electricity on an industrial scale.



**Figure 6.11** The nuclear power complex at Sellafield, Cumbria, UK was the first to produce electricity in the world.

Nuclear reactors harness the energy from the fission of uranium-235. **Nuclear fission** occurs when the unstable nucleus of a radioactive isotope splits up, forming smaller atoms and producing a large amount of energy as a result. Scientists believe that the energy comes from the conversion of some of the mass of the isotope.

This fission process begins when a neutron hits an atom of uranium-235, causing it to split and produce



**Figure 6.12** Chain reaction in uranium-235 fission.

three further neutrons. These three neutrons split three more atoms of uranium-235, which produces nine neutrons and so on. This initiates a **chain reaction** (Figure 6.12).

In a reactor the fission process cannot be allowed to get out of control as it does in an atomic bomb. To prevent this, boron control rods can be pushed into different positions in the reactor to absorb some of the neutrons which are produced and so slow down the chain reaction. If this is done, the energy released from the reaction is obtained in a more controlled way. The energy is used to produce steam, which in turn is used to generate electricity (Figure 6.10, p. 92).

However, there are problems. The main problem associated with a nuclear power station is that the reactor produces highly radioactive waste materials. These waste materials are difficult to store and cannot be disposed of very easily. Also, leaks of radioactive material have occurred at various sites throughout the world. Accidents at a small number of nuclear power stations, such as Three Mile Island in the US (1979) and Chernobyl in the Ukraine (1986) have led to a great deal of concern about their safety. More recently, in March 2011 a major nuclear accident happened at the Fukushima nuclear power plant in northern Japan. On this occasion the accident was not caused as a result of the plant itself undergoing a problem, but as a result of an earthquake near Japan that gave rise to a tsunami. This damaged essential

equipment at the plant which was used to ensure the temperatures were maintained at a correct level. Work is still going on at the plant to remove the nuclear material safely.



**Figure 6.13** Workers wearing protective suits and masks at the Fukushima power plant.

## Hydroelectric power

Hydroelectric power (HEP) is electricity generated from the energy of falling water (Figure 6.14). It is an excellent energy source, and electricity has been generated in this way for some time. For example, the Hoover Dam in the US provides HEP for the surrounding area. It is a very cheap source of electricity. Once you have built the power station, the energy is absolutely free. In some mountainous areas of the world, such as the Alps, HEP is the main source of electricity. One of the main advantages of this system is that it can be quickly used to supplement the national grid at times of high demand. A disadvantage of HEP schemes is that they often require valleys to be flooded and communities to be moved.



**Figure 6.14** A hydroelectric power station.

## Biomass and biogas

When any biological material, whether plant or animal, is converted into energy, this energy is called **biomass** energy. It can be taken from animal or plant materials in different ways:

- by burning it, for example wood (Figure 6.15)



**Figure 6.15** Biomass energy is produced by burning wood.

- by pressing out oils that can be burned
- by fermenting it to produce fuels such as ethanol or methane.

At least 50% of the world's population rely on wood as their main energy source.

In India there are millions of methane generators. Methane generated by the digestion of animal waste is called **biogas**. The biogas produced is used for cooking, heating and lighting. The by-product of this process is an excellent fertiliser.

Some countries have already experimented with ethanol as a fuel for cars. Up to 20% of ethanol can be added to petrol without the need to adjust the carburettor. Brazil, which has few oil reserves, produces ethanol by fermentation (breakdown by enzymes) of sugar cane and grain, and uses it as a petrol additive (Figure 6.16). The Brazilian government has cut down its petrol imports by up to 60% through using this alcohol/petrol mixture.



**Figure 6.16** In Brazil cars use an ethanol/petrol mixture.

## Hydrogen – the fuel of the future

It is widely believed that hydrogen will play an important part as a fuel in the future. It is already being trialled by major motor manufacturers as an alternative to fossil fuels such as petrol (Figure 6.17).

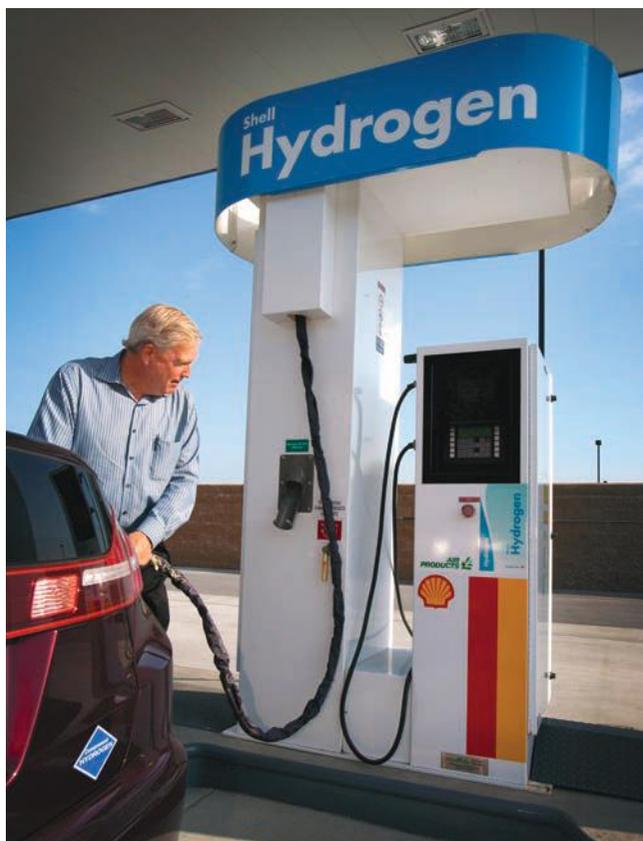
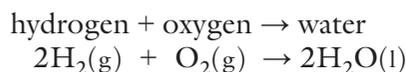


Figure 6.17 This car is powered by hydrogen gas.

Hydrogen burns cleanly to form water:



Hence virtually no exhaust pollution is created. Also, hydrogen can be obtained in an almost unlimited supply from water by electrolysis. However, at present it is a costly alternative.

### Questions

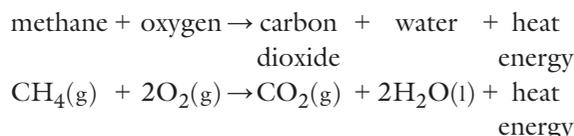
- 1 Draw up a table showing the alternative sources of energy along with their advantages and disadvantages.
- 2 What is meant by the terms:
  - a non-renewable energy sources?
  - b renewable energy sources?
- 3 Use your research skills to obtain technical information about the operation of Pelamis Wave Energy Converter.

## Chemical energy

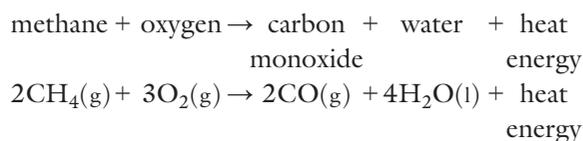
We obtain our energy needs from the combustion of fuels, such as hydrocarbons, from the combustion of foods and from many other chemical reactions.

### Combustion

When natural gas burns in a plentiful supply of air it produces a large amount of energy.



During this process, the **complete combustion** of methane, heat is given out. It is an **exothermic** reaction. If only a limited supply of air is available then the reaction is not as exothermic and the poisonous gas carbon monoxide is produced.



This is the **incomplete combustion** of methane.

The energy changes that take place during a chemical reaction can be shown by an **energy level diagram**. Figure 6.18 shows the energy level diagram for the complete combustion of methane.

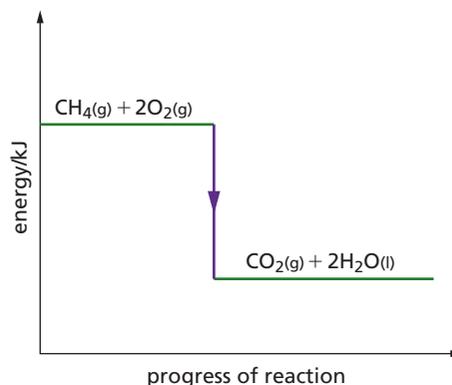


Figure 6.18 Energy level diagram for the complete combustion of methane.

When any reaction occurs, the chemical bonds in the reactants have to be broken – this requires energy. This is an endothermic process. When the new bonds in the products are formed, energy is given out (Figure 6.19). This is an exothermic process. The **bond energy** is defined as the amount of energy in kilojoules (kJ) associated with the breaking or making of one mole of chemical bonds in a molecular element or compound.

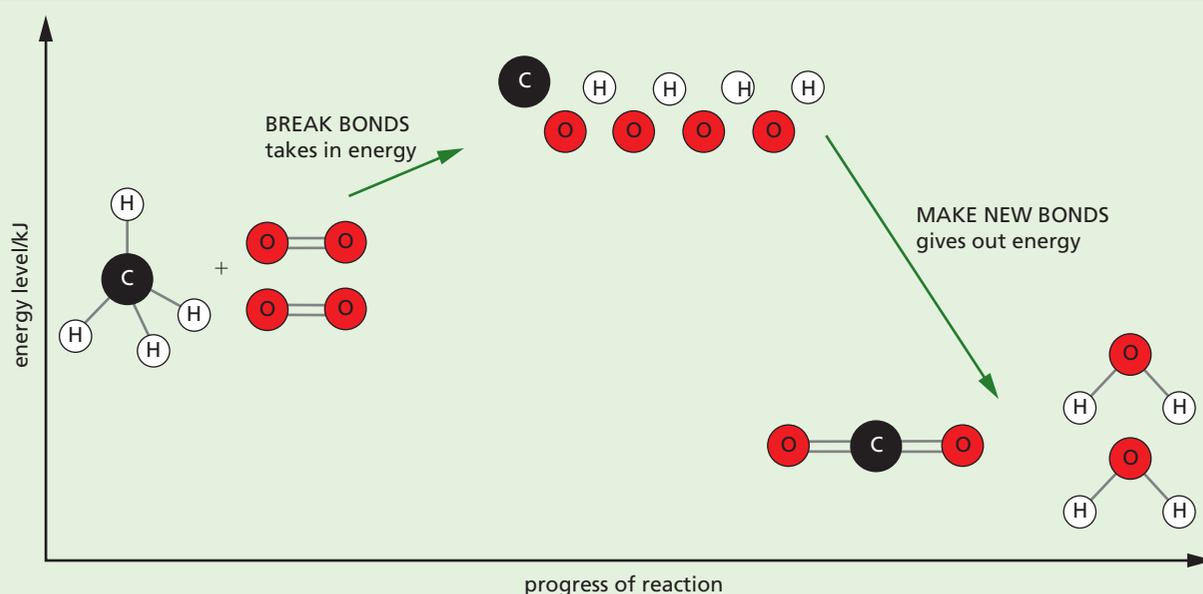


Figure 6.19 Breaking and forming bonds during the combustion of methane.

Using the data in Table 6.3, which tells us how much energy is needed to break a chemical bond and how much is given out when it forms, we can calculate how much energy is involved in each stage.

Table 6.3 Bond energy data.

Bond	Bond energy/kJ mol <sup>-1</sup>
C—H	435
O=O	497
C=O	803
H—O	464
C—C	347
C—O	358

### Bond breaking

Breaking 4 C—H bonds in methane requires

$$4 \times 435 = 1740 \text{ kJ}$$

Breaking 2 O=O bonds in oxygen requires

$$2 \times 497 = 994 \text{ kJ}$$

Total = 2734 kJ of energy

### Making bonds

Making 2 C=O bonds in carbon dioxide gives out

$$2 \times 803 = 1606 \text{ kJ}$$

Making 4 O—H bonds in water gives out

$$4 \times 464 = 1856 \text{ kJ}$$

Total = 3462 kJ of energy

Energy difference

$$= \text{energy required to break bonds} - \text{energy given out when bonds are made}$$

$$= 2734 - 3462$$

$$= -728 \text{ kJ}$$

The negative sign shows that the chemicals are losing energy to the surroundings, that is, it is an exothermic reaction. A positive sign would indicate that the chemicals are gaining energy from the surroundings. This type of reaction is called an **endothermic** reaction.

The energy stored in the bonds is called the **enthalpy** and is given the symbol  $H$ . The change in energy going from reactants to products is called the **change in enthalpy** and is shown as  $\Delta H$  (pronounced 'delta H').  $\Delta H$  is called the **heat of reaction**.

For an exothermic reaction  $\Delta H$  is negative and for an endothermic reaction  $\Delta H$  is positive.

When fuels, such as methane, are burned they require energy to start the chemical reaction. This is known as the **activation energy**,  $E_A$  (Figure 6.20). In the case of methane reacting with oxygen, it is some of the energy involved in the initial bond breaking (Figure 6.20). The value of the activation energy will vary from fuel to fuel.

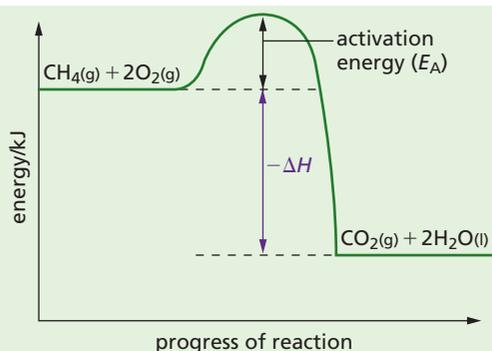


Figure 6.20 Energy level diagram for methane/oxygen.

Endothermic reactions are much less common than exothermic ones. In this type of reaction energy is absorbed from the surroundings so that the energy of the products is greater than that of the reactants. The reaction between nitrogen and oxygen gases is endothermic (Figure 6.21), and the reaction is favoured by high temperatures.

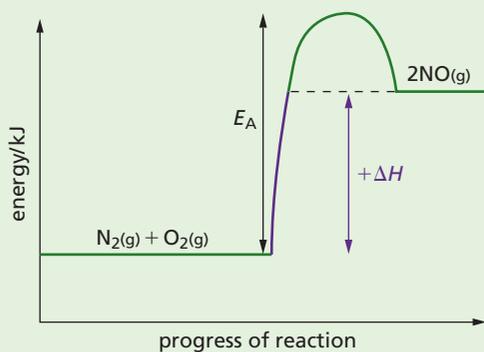
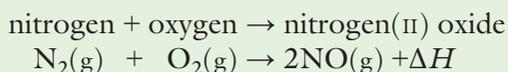
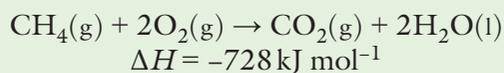


Figure 6.21 Energy level diagram for nitrogen/oxygen.

Dissolving is often an endothermic process. For example, when ammonium nitrate dissolves in water the temperature of the water falls, indicating that energy is being taken from the surroundings. Photosynthesis and thermal decomposition are other examples of endothermic processes.

In equations it is usual to express the  $\Delta H$  value in units of  $\text{kJ mol}^{-1}$ . For example:



This  $\Delta H$  value tells us that when 1 mole of methane is burned in oxygen, 728 kJ of energy are released. This value is called the **enthalpy of combustion** (or **molar heat of combustion**) of methane.

## Enthalpy of neutralisation (molar heat of neutralisation)

This is the enthalpy change that takes place when 1 mol of hydrogen ions ( $\text{H}^+(\text{aq})$ ) is neutralised.



This process occurs in the titration of an alkali by an acid to produce a neutral solution (Chapter 8, p. 124).

## Questions

- Using the bond energy data given in Table 6.3:
  - Calculate the enthalpy of combustion of ethanol, a fuel added to petrol in some countries.
  - Draw an energy level diagram to represent this combustion process.
  - How does this compare with the enthalpy of combustion of heptane ( $\text{C}_7\text{H}_{14}$ ), a major component of petrol, of  $-4853 \text{ kJ mol}^{-1}$ ?
  - How much energy is released per gram of ethanol and heptane burned?
- How much energy is released if:
  - 0.5 mole of methane is burned?
  - 5 moles of methane are burned?
  - 4 g of methane are burned?  
( $A_r$ : C = 12; H = 1)
- How much energy is released if:
  - 2 moles of hydrogen ions are neutralised?
  - 0.25 mole of hydrogen ions is neutralised?
  - 1 mole of sulfuric acid is completely neutralised?

## Changes of state

In Chapter 1, p. 4, we discussed the melting and boiling of a substance. The heating curve for water is shown in Figure 1.11 on p. 5. For ice to melt to produce liquid water, it must absorb energy from its surroundings. This energy is used to break down the weak forces between the water molecules (intermolecular forces) in the ice. This energy is called the **enthalpy of fusion** and is given the symbol  $\Delta H_{\text{fusion}}$ . Similarly, when liquid water changes into steam, the energy required for this process to occur is called the **enthalpy of vaporisation** and is given the symbol  $\Delta H_{\text{vap}}$ . Figure 6.22 shows an energy level diagram representing both the fusion and the vaporisation processes.

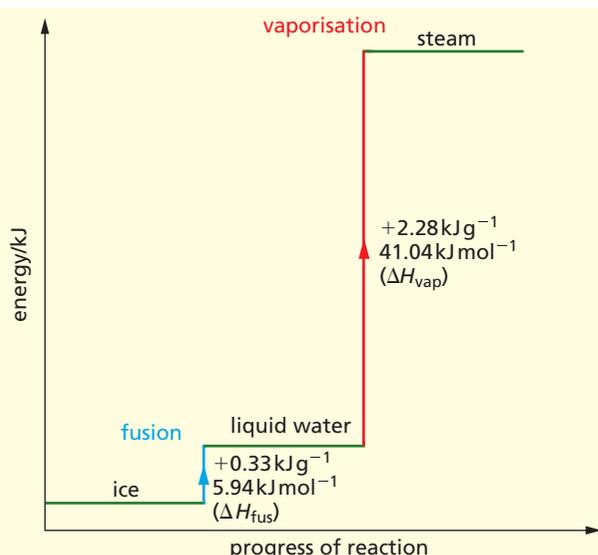


Figure 6.22 Energy level diagram for the fusion and vaporisation of water.

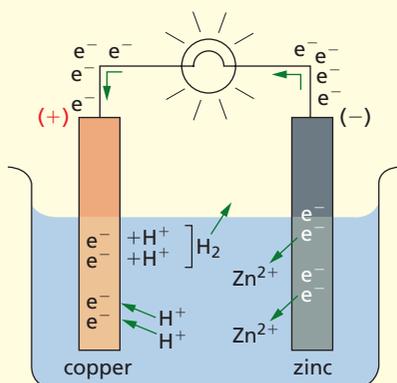
### Questions

- Describe the energy changes which take place when the processes described in this section, with water, are reversed.
- Using the knowledge you have obtained from Chapter 1, p. 5, give a full definition of the enthalpy of fusion and enthalpy of vaporisation for water.

## Cells and batteries

A chemical cell is one which produces electrical energy from chemical energy. The electrical energy is produced as a result of the transfer of electrons which takes place during the chemical reaction via a redox process.

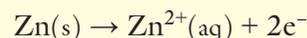
A simple type of chemical cell is that shown in Figure 6.23a. In this cell the more reactive metal



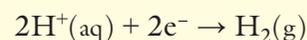
a A simple chemical cell.

Figure 6.23

zinc dissolves in the dilute sulfuric acid, producing zinc ions ( $\text{Zn}^{2+}(\text{aq})$ ) and releasing two electrons.

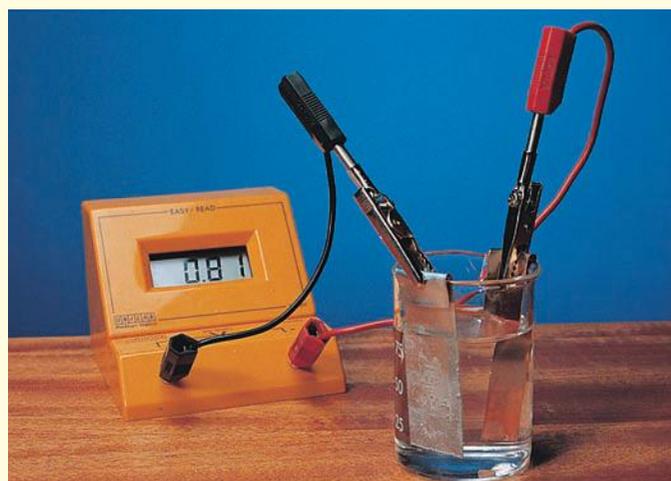


These electrons produced at the zinc electrode cause it to become the negative terminal. They flow to the positive copper terminal through the external circuit via the bulb, which glows. Bubbles of hydrogen are seen when the electrons arrive at the copper electrode. The hydrogen gas is produced from the hydrogen ions in the acid, which collect the electrons appearing at the copper electrode.



Slowly, the zinc electrode dissolves in the acid and the bulb will then go out. If the zinc is replaced by a more reactive metal, such as magnesium, then the bulb glows more brightly. Magnesium loses electrons more easily as it reacts faster with the dilute acid.

The difference in the reactivity between the two metals used in the cell creates a particular voltage reading on the voltmeter shown in Figure 6.23b. The more the two metals differ in reactivity, the larger is the voltage shown and delivered by the cell. This method can be used to confirm the order of reactivity of the metals (Chapter 10, p. 150). Other types of chemical cell in common use are dry cells used in radios, torches, and so on, and lead–acid accumulators used in motor vehicles. These are convenient and portable energy sources.



b The voltage reflects the difference in reactivity of the metals.

## Fuel cells

Scientists have found a much more efficient way of changing chemical energy into electrical energy, using a fuel cell (Figure 6.24). Fuel cells are like the chemical cells in the previous section, except that the reagents are supplied continuously to the electrodes. The reagents are usually hydrogen and oxygen. The fuel cell principle was first discovered by Sir William Grove in 1839.

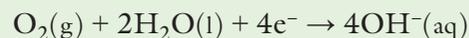


**Figure 6.24** The space shuttle's computers use electricity produced by fuel cells.

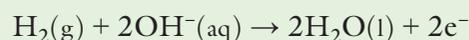
When he was electrolysing water and he switched off the power supply, he noticed that a current still flowed but in the reverse direction. Subsequently, the process was explained in terms of the reactions at the electrodes' surfaces of the oxygen and hydrogen gases which had been produced during the electrolysis.

The hydrogen fuel cells used by NASA in the US space programme are about 70% efficient and, since the only product is water, they are pollution free. The aqueous NaOH electrolyte is kept within the cell by electrodes which are porous, allowing

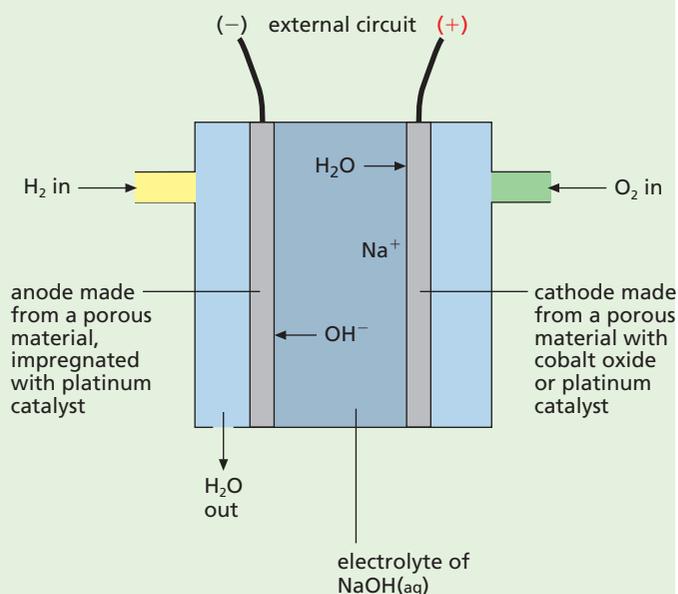
the transfer of  $O_2$ ,  $H_2$  and water through them (Figure 6.25). As  $O_2$  gas is passed into the cathode region of the cell it is reduced:



The  $OH^-$  ions formed are removed from the fuel cell by reaction with  $H_2$ :



The electrons produced by this process pass around an external circuit to the cathode.



**Figure 6.25** A diagrammatic view of a fuel cell.

## Questions

- 1 Describe how simple chemical cells can be used to confirm the order of reactivity of the metals in the reactivity series.
- 2 The fuel cell was discovered during electrolysis experiments with water. It is the reverse process which produces the electricity. Write a balanced chemical equation to represent the overall reaction taking place in a fuel cell.

### Checklist

After studying Chapter 6 you should know and understand the following terms.

- **Aerobic decay** Decay which takes place in the presence of oxygen.
- **Anaerobic decay** Decay which takes place in the absence of oxygen.
- **Bond energy** An amount of energy associated with a particular bond in a molecular element or compound.
- **Chain reaction** A nuclear reaction which is self-sustaining as a result of one of the products causing further reactions.
- **Chemical cell** A system for converting chemical energy to electrical energy.
- **Combustion** A chemical reaction in which a substance reacts rapidly with oxygen with the production of heat and light.
- **Endothermic reaction** A chemical reaction which absorbs heat energy from its surroundings.
- **Enthalpy** Energy stored in chemical bonds, given the symbol  $H$ .
- **Enthalpy change** Given the symbol  $\Delta H$ , it represents the difference between energies of reactants and products.
- **Enthalpy of combustion** The enthalpy change which takes place when one mole of a substance is completely burned in oxygen.
- **Enthalpy of fusion** The enthalpy change that takes place when one mole of a solid is changed to one mole of liquid at the same temperature.
- **Enthalpy of neutralisation** The enthalpy change which takes place when one mole of hydrogen ions is completely neutralised.
- **Enthalpy of vaporisation** The enthalpy change that takes place when one mole of liquid is changed to one mole of vapour at the same temperature.
- **Exothermic reaction** A chemical reaction that releases heat energy into its surroundings.
- **Fossil fuels** Fuels, such as coal, oil and natural gas, formed from the remains of plants and animals.
- **Hydrocarbon** A substance which contains atoms of carbon and hydrogen only.
- **Non-renewable energy sources** Sources of energy, such as fossil fuels, which take millions of years to form and which we are using up at a rapid rate.
- **Nuclear fission** The disintegration of a radioactive nucleus into two or more lighter fragments. The energy released in the process is called nuclear energy.
- **Oil refining** The general process of converting the mixture that is collected as crude oil into separate fractions. These fractions, known as petroleum products, are used as fuels, lubricants, bitumens and waxes. The fractions are separated from the crude oil mixture by fractional distillation.
- **Organic compounds** Substances whose molecules contain one or more carbon atoms covalently bonded with another element (including hydrogen, nitrogen, oxygen, the halogens as well as phosphorus, silicon and sulfur).
- **Renewable energy** Sources of energy which cannot be used up or which can be made at a rate faster than the rate of use.

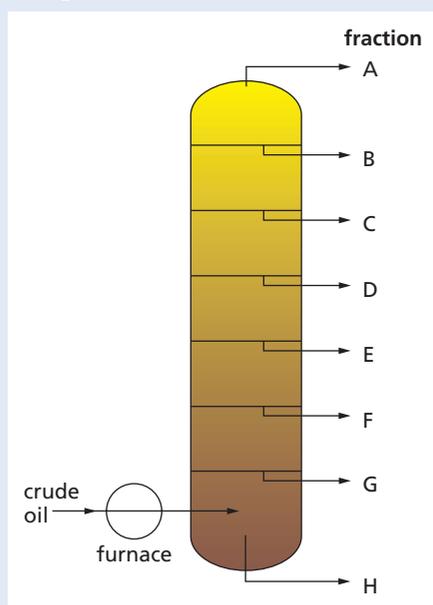
# Chemical energetics

## Additional questions

- 1 a State which of the following processes is endothermic and which is exothermic.
- The breaking of a chemical bond.
  - The forming of a chemical bond.
- b The table below shows the bond energy data for a series of covalent bonds.
- Use the information given in the table to calculate the overall enthalpy change for the combustion of ethanol producing carbon dioxide and water.
  - Is the process in (i) endothermic or exothermic?

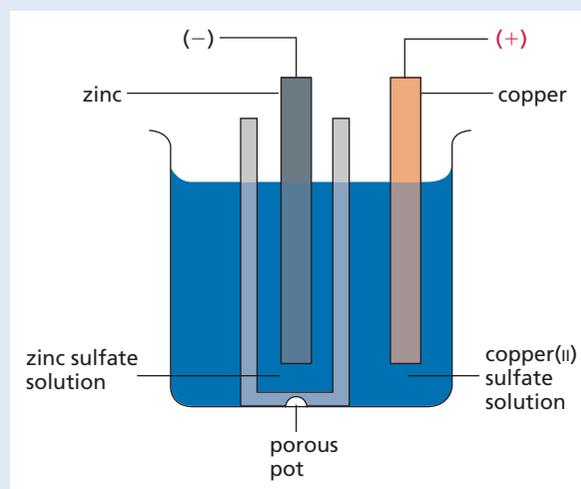
Bond	Bond energy/ kJ mol <sup>-1</sup>
C—H	435
O=O	497
C=O	803
H—O	464
C—C	347
C—O	358

- 2 Crude oil is a mixture of hydrocarbons. The refining of crude oil produces fractions which are more useful to us than crude oil itself. Each fraction is composed of hydrocarbons which have boiling points within a specific range of temperature. The separation is carried out in a fractionating column, as shown below.



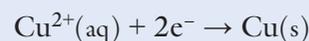
- a Which separation technique is used to separate the fractions?

- b Name each of the fractions A to H and give a use for each.
- c Why do the fractions come from the fractionating column in this order?
- d What is the connection between your answer to c and the size of the molecules in each fraction?
- e Which of the fractions will be the most flammable?
- 3 Explain the following.
- Hydroelectric power is a relatively cheap source of electricity.
  - Geothermal energy is a non-polluting form of energy.
  - A disadvantage of wind power is that it causes noise pollution.
  - The by-product from the process by which methane is generated by the digestion of animal waste is an excellent fertiliser.
  - The fission of uranium-235 in a nuclear reactor is an example of a chain reaction.
  - Tidal- and wave-generated electricity has a major environmental disadvantage.
- 4 One of the first practical chemical cells was the Daniell cell invented by John Daniell in 1836. A diagram of this type of cell is shown below.



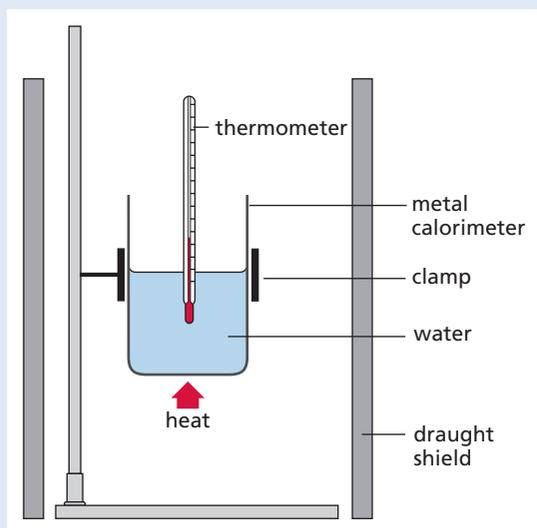
It is capable of generating about 1.1 volts and was used to operate small electrical items such as doorbells.

- a The electrode reaction taking place at a copper anode is:



Write an electrode equation for the process taking place at the cathode.

- b Which way would the electrons flow in the wire connected to the voltmeter – from ‘copper to zinc’ or ‘zinc to copper’?
- c Why should copper(II) sulfate crystallise at the bottom of the outer container?
- d What is the function of the porous pot?
- e There are problems associated with the Daniell cell which have led to it being replaced by other types of cell. Give two reasons why Daniell cells are no longer in use today.
- 5 This question is about endothermic and exothermic reactions.
- a Explain the meaning of the terms endothermic and exothermic.
- b (i) Draw an energy level diagram for the reaction:
- $$\text{NaOH(aq)} + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$$
- $$\Delta H = -57 \text{ kJ mol}^{-1}$$
- (ii) Is this reaction endothermic or exothermic?
- (iii) Calculate the energy change associated with this reaction if 2 moles of sodium hydroxide were neutralised by excess hydrochloric acid.
- c (i) Draw an energy level diagram for the reaction:
- $$2\text{H}_2\text{O(l)} \rightarrow 2\text{H}_2\text{(g)} + \text{O}_2\text{(g)}$$
- $$\Delta H = +575 \text{ kJ mol}^{-1}$$
- (ii) Is this reaction endothermic or exothermic?
- (iii) Calculate the energy change for this reaction if only 9 g of water were converted into hydrogen and oxygen.
- 6 The following results were obtained from an experiment carried out to measure the enthalpy of combustion (heat of combustion) of ethanol. The experiment involved heating a known volume of water with the flame from an ethanol burner.



The burner was weighed initially and after the desired temperature rise had been obtained.

Volume of water in glass beaker = 200 cm<sup>3</sup>

Mass of ethanol burner at start = 85.3 g

Mass of ethanol burner at end = 84.8 g

Temperature rise of water = 12 °C

(Density of water = 1 g cm<sup>-3</sup>)

Heat energy = mass of water/g × 4.2 J g<sup>-1</sup> °C<sup>-1</sup> × temperature rise/°C

- a Calculate the mass of ethanol burned.
- b Calculate the amount of heat produced, in joules, in this experiment by the ethanol burning.
- c Convert your answer to b into kilojoules.
- d Calculate the amount of heat produced by 1 g of ethanol burning.
- e What is the mass of 1 mole of ethanol (C<sub>2</sub>H<sub>5</sub>OH)?  
(A<sub>r</sub>: H = 1; C = 12; O = 16)
- f How much heat would be produced if 1 mole of ethanol had been burned? (This is the heat of combustion of ethanol.)
- g Compare your value with the actual value of 1371 kJ mol<sup>-1</sup> and suggest two reasons for the difference in values.
- h Write a balanced chemical equation to represent the combustion of ethanol.
- 7 The following results were obtained from a neutralisation reaction between 1 mol dm<sup>-3</sup> hydrochloric acid and 1 mol dm<sup>-3</sup> sodium hydroxide. This experiment was carried out to measure the heat of neutralisation of hydrochloric acid. The temperature rise which occurred during the reaction was recorded.

Volume of sodium hydroxide used = 50 cm<sup>3</sup>

Volume of acid used = 50 cm<sup>3</sup>

Temperature rise = 5 °C

(Density of water = 1 g cm<sup>-3</sup>)

Heat energy = mass of water/g × 4.2 J g<sup>-1</sup> °C<sup>-1</sup> × temperature rise/°C

- a Write a balanced chemical equation for the reaction.
- b What mass of solution was warmed during the reaction?

- c How much heat energy was produced during the reaction?
- d How many moles of hydrochloric acid were involved in the reaction?
- e How much heat would be produced if 1 mole of hydrochloric acid had reacted? (This is the heat of neutralisation of hydrochloric acid.)
- f The heat of neutralisation of hydrochloric acid is  $-57 \text{ kJ mol}^{-1}$ . Suggest two reasons why there is a difference between this and your calculated value.
- 8 Write down which factors are most important when deciding on a particular fuel for the purpose given:
- fuel for a cigarette lighter
  - fuel for a camping stove
  - fuel for an aeroplane
  - fuel for an underground transport system
  - fuel for a space craft
  - fuel for domestic heating.
- 9 'Propagas' is used in some central heating systems where natural gas is not available. It burns according to the following equation:
- $$\text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{l})$$
- $$\Delta H = -2220 \text{ kJ mol}^{-1}$$
- What are the chemical names for 'propagas' and natural gas?
  - Would you expect the heat generated per mole of 'propagas' burned to be greater than that for natural gas? Explain your answer.
  - What is 'propagas' obtained from?
  - Calculate:
    - the mass of 'propagas' required to produce 5550 kJ of energy
    - the heat energy produced by burning 0.5 mole of 'propagas'
    - the heat energy produced by burning 11 g of 'propagas'
    - the heat energy produced by burning 2000 dm<sup>3</sup> of 'propagas'.( $A_r$ : H = 1; C = 12; O = 16. One mole of any gas occupies 24 dm<sup>3</sup> at room temperature and pressure.)

## 7

# Chemical reactions

**Factors that affect the rate of a reaction**

Surface area  
Concentration  
Temperature  
Light  
Catalysts

**Enzymes****Checklist****Additional questions**

Figure 7.1 shows some slow and fast reactions. The two photographs on the left show examples of slow reactions. The ripening of apples takes place over a number of weeks, and the making and maturing of cheese may take months. The burning of solid fuels, such as coal, can be said to involve chemical reactions taking place at a medium speed or rate. The other example shows a fast reaction. The chemicals inside explosives, such as TNT, react very rapidly in reactions which are over in seconds or fractions of seconds.

As new techniques have been developed, the processes used within the chemical industry have become more complex. Therefore, chemists and chemical engineers have increasingly looked for ways to control the rates at which chemical reactions take place. In doing so, they have discovered that there are five main ways in which you can alter the rate of a chemical reaction. These ideas are not only incredibly useful to industry but can also be applied to reactions which occur in the school laboratory.



**Figure 7.1** Some slow (ripening fruit and cheese making), medium (coal fire) and fast (explosion) reactions.

## ● Factors that affect the rate of a reaction

- Surface area of the reactants.
- Concentration of the reactants and gas pressure.
- Temperature at which the reaction is carried out.
- Light.
- Use of a catalyst, including enzymes.

### Collision theory

For a chemical reaction to occur, reactant particles need to collide with one another. Not every collision results in the formation of products. For products to be formed, the collision has to have a certain minimum amount of energy associated with it. This minimum amount of energy is known as the **activation energy**,  $E_a$  (Figure 7.2). Collisions which result in the formation of products are known as successful collisions.

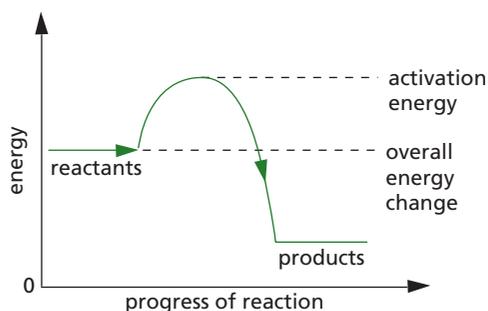


Figure 7.2 Energy level diagram showing activation energy.

### Surface area

In Chapter 13, we shall see that limestone (calcium carbonate) is a substance which can be used to neutralise soil acidity. Powdered limestone is used as it neutralises the acidity faster than if lumps of limestone are used. Why do you think this is the case?

In the laboratory, the reaction between acid and limestone in the form of lumps or powder can be observed in a simple test-tube experiment. Figure 7.3 shows the reaction between dilute hydrochloric acid and limestone in lump and powdered form.

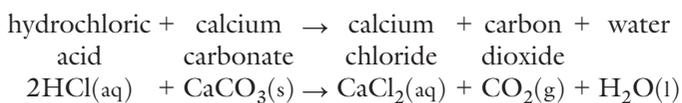


Figure 7.3 The powdered limestone (left) reacts faster with the acid than the limestone in the form of lumps.

The rates at which the two reactions occur can be found by measuring either:

- the volume of the carbon dioxide gas which is produced, or
- the loss in mass of the reaction mixture with time.

These two methods are generally used for measuring the rate of reaction for processes involving the formation of a gas as one of the products.

The apparatus shown in Figure 7.4 (p. 106) is used to measure the loss in mass of the reaction mixture. The mass of the conical flask plus the reaction mixture is measured at regular intervals. The total loss in mass is calculated for each reading of the balance, and this is plotted against time. Some sample results from experiments of this kind have been plotted in Figure 7.5.

The reaction is generally at its fastest in the first minute. This is indicated by the slopes of the curves during this time. The steeper the slope, the faster the rate of reaction. You can see from the two traces in Figure 7.5 that the rate of reaction is greater with the powdered limestone than the lump form.

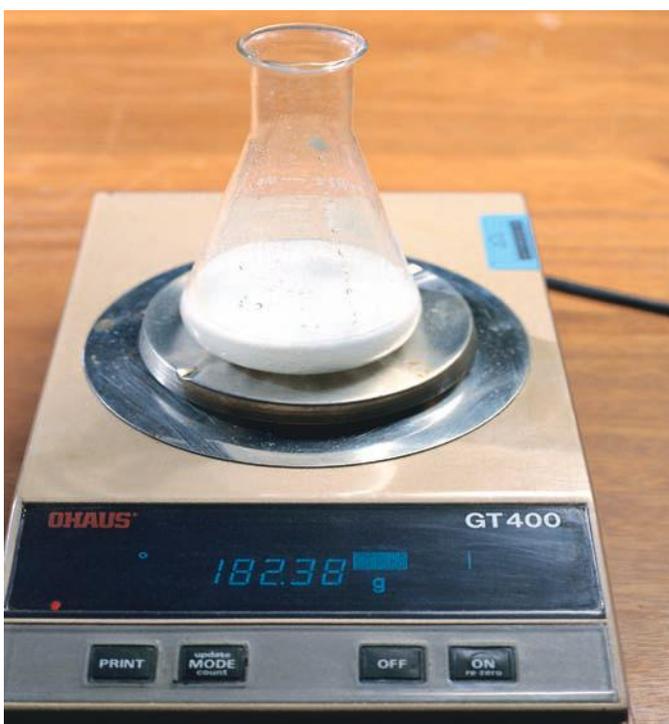
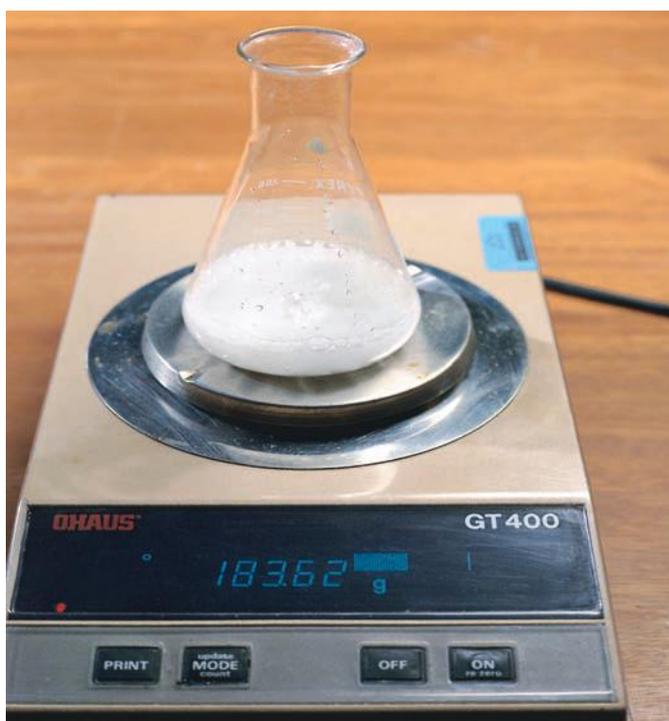


Figure 7.4 After 60 seconds the mass has fallen by 1.24 g.

The surface area has been increased by powdering the limestone (Figure 7.6). The acid particles now have an increased amount of surface of limestone with which to collide. The products of a reaction are formed when collisions occur between reactant particles.

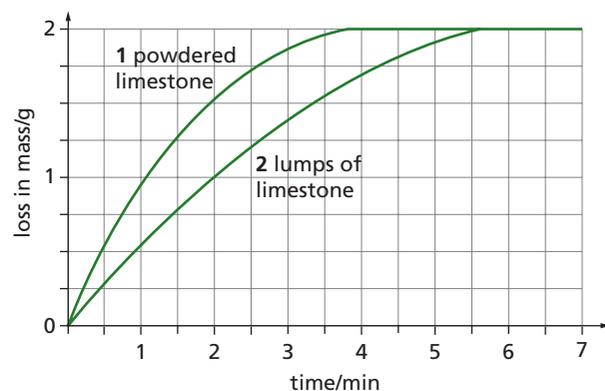


Figure 7.5 Sample results for the limestone/acid experiment.

An increase in the surface area of a solid reactant results in an increase in the number of collisions, and this results in an increase in the number of successful collisions. Therefore, the increase in surface area of the limestone increases the rate of reaction.

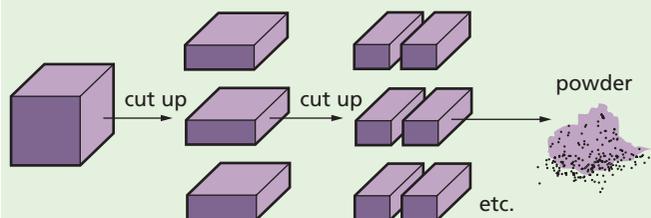


Figure 7.6 A powder has a larger surface area.

In certain industries the large surface area of fine powders and dusts can be a problem. For example, there is a risk of explosion in flourmills and mines, where the large surface area of the flour or coal dust can – and has – resulted in explosions through a reaction with oxygen gas in the air when a spark has been created by machinery or the workforce (Figure 7.7). On 26 September 1988, two silos containing wheat exploded at the Jamaica Flour Mills Plant in Kingston, Jamaica, killing three workers, as a result of fine dust exploding.



Figure 7.7 The dust created by this cement plant in India is a potential hazard.

## Questions

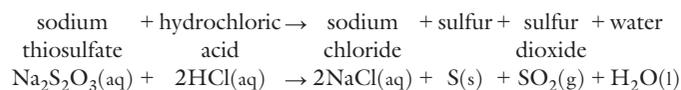
- 1 What apparatus would you use to measure the rate of reaction of limestone with dilute hydrochloric acid by measuring the volume of carbon dioxide produced?
- 2 The following results were obtained from an experiment of the type you were asked to design in question 1.

Time/min	0	0.5	1.0	1.5	2.0	2.5	3.0	3.5	4.0	4.5	5.0
Total volume of CO <sub>2</sub> gas/cm <sup>3</sup>	0	15	24	28	31	33	35	35	35	35	35

- a Plot a graph of the total volume of CO<sub>2</sub> against time.
- b At which point is the rate of reaction fastest?
- c What volume of CO<sub>2</sub> was produced after 1 minute 15 seconds?
- d How long did it take to produce 30 cm<sup>3</sup> of CO<sub>2</sub>?

## Concentration

A yellow precipitate is produced in the reaction between sodium thiosulfate and hydrochloric acid.



The rate of this reaction can be followed by recording the time taken for a given amount of sulfur to be precipitated. This can be done by placing a conical flask containing the reaction mixture on to a cross on a piece of paper (Figure 7.8). As the precipitate of sulfur forms, the cross is obscured and finally disappears from view. The time taken for this to occur is a measure of the rate of this reaction. To obtain sufficient information about the effect of changing the concentration of the reactants, several experiments of this type must be carried out, using different concentrations of sodium thiosulfate or hydrochloric acid.

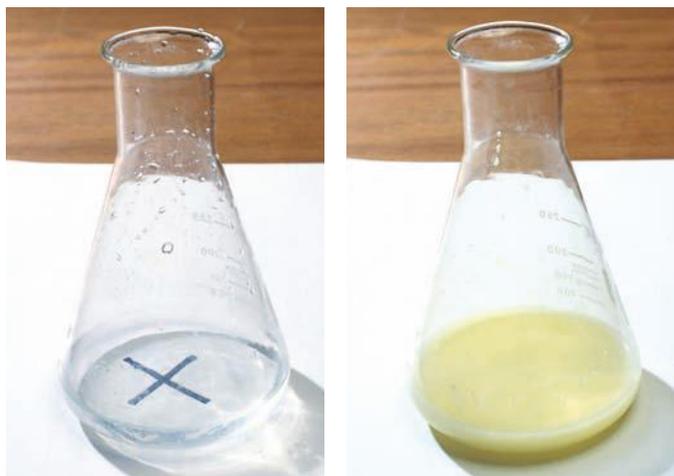


Figure 7.8 The precipitate of sulfur obscures the cross.

Some sample results of experiments of this kind have been plotted in Figure 7.9. You will note from the graph that when the most concentrated sodium thiosulfate solution was used, the reaction was at its fastest. This is shown by the shortest time taken for the cross to be obscured.

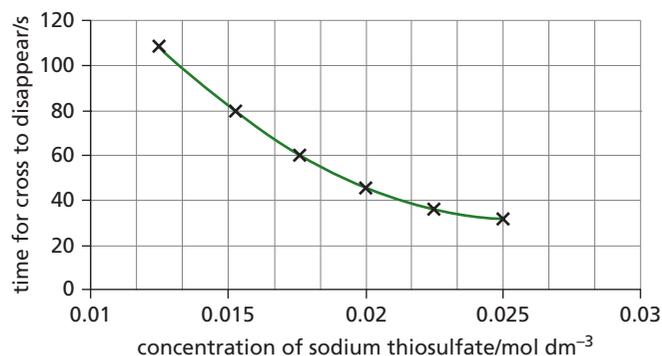


Figure 7.9 Sample data for the sodium thiosulfate/acid experiment at different concentrations of sodium thiosulfate.

From the data shown in Figure 7.9 it is possible to produce a different graph which directly shows the rate of the reaction against concentration rather than the time taken for the reaction to occur against concentration. To do this, the times can be converted to a rate using:

$$\text{rate} = \frac{1}{\text{reaction time (s)}}$$

This would give the graph shown in Figure 7.10 (p. 108).

As discussed earlier, the products of the reaction are formed as a result of the collisions between reactant particles. There are more particles in a more concentrated solution and the collision rate between reactant particles is higher. The more often the particles collide, the greater the chance they have of having sufficient energy to overcome the activation energy of the reaction, and of a successful collision occurring. This means that the rate of a chemical reaction will increase if the concentration of reactants is increased, because there are more particles per unit volume.

In reactions involving only gases, for example the Haber process (Chapter 11, p. 177), an increase in the overall pressure at which the reaction is carried out increases the rate of the reaction. The increase in pressure results in the gas particles being pushed closer together. This means that they collide more often and so react faster.

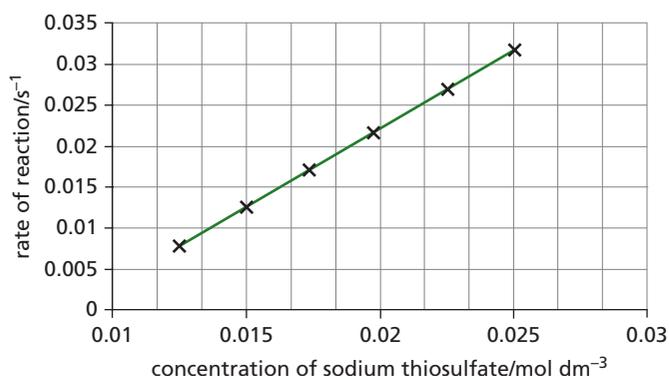


Figure 7.10 Graph to show the rate of reaction against concentration.

### Question

- 1 Devise an experiment to show the effect of changing the concentration of dilute acid on the rate of reaction between magnesium and hydrochloric acid.

## Temperature

Why do you think food is stored in a refrigerator? The reason is that the rate of decay is slower at lower temperatures. This is a general feature of the majority of chemical processes.

The reaction between sodium thiosulfate and hydrochloric acid can also be used to study the effect of temperature on the rate of a reaction. Figure 7.11 shows some sample results of experiments with sodium thiosulfate and hydrochloric acid (at fixed concentrations) carried out at different temperatures. You can see from the graph that the rate of the reaction is fastest at high temperatures.

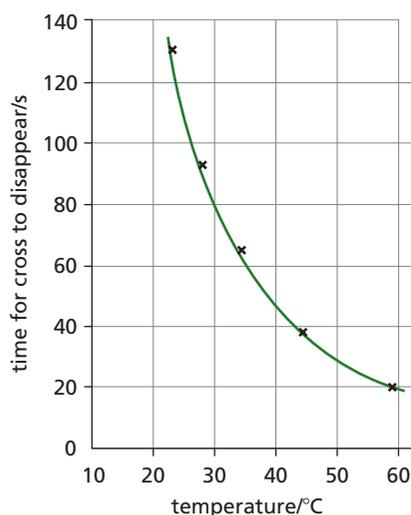


Figure 7.11 Sample data for the sodium thiosulfate/acid experiment at different temperatures.

As the temperature increases, the reactant particles increase their kinetic energy and they move faster. The faster movement results in more collisions between the particles. Some of these extra collisions, which result from the temperature increase, will be successful collisions. This causes the reaction rate to increase.

### Questions

- 1 Explain why potatoes cooked in oil cook faster than those cooked in water.
- 2 Devise an experiment to study the effect of temperature on the reaction between magnesium and hydrochloric acid.
- 3 Explain why food cooks faster in a pressure cooker.

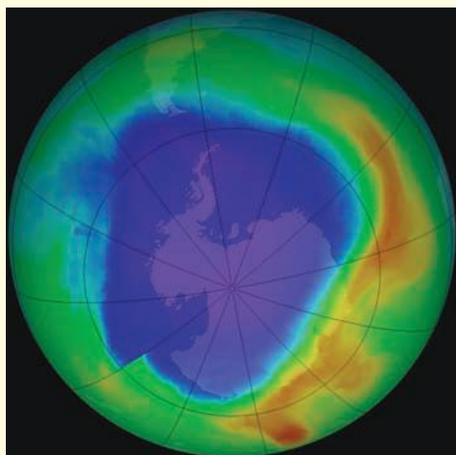
## Light

Some chemical reactions are affected by light. When particles absorb light energy, the energy can be used to break bonds, overcoming the activation energy of the reactions and causing a chemical reaction to occur faster (see p. 105). Reactions which occur as a result of the absorption of light are known as **photochemical reactions**. The absorption of light in these reactions causes bonds to break, producing reactive particles known as radicals. These radicals are responsible for many of the chemical reactions which happen in the stratosphere (see p. 221).

### Ozone depletion

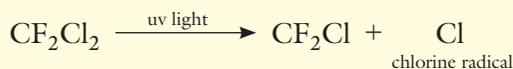
Ozone, O<sub>3</sub>, is an important gas found in the stratosphere. It is important because it absorbs high-energy ultraviolet radiation. If this radiation were to reach the surface of the Earth it would lead to an increase in skin cancers and other medical problems.

In 1985 large holes were discovered in the ozone layer over Antarctica, Australasia and Europe (Figure 7.12). Scientists think that these holes have partly been produced by chemicals called chlorofluorocarbons or CFCs. CFCs were developed as refrigerants in fridges back in the 1950s. They have also been used in air conditioning systems, in aerosol cans (as propellants) and in the manufacture of expanded plastics such as polystyrene. CFCs are very unreactive molecules. They escape into the atmosphere and, because of their inertness, remain without further reaction until they reach the

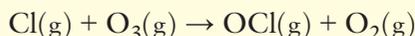


**Figure 7.12** Diagram showing the hole in the ozone layer over Antarctica

stratosphere and the ozone layer. In the stratosphere, the high-energy ultraviolet radiation causes a chemical bond in the CFC molecule to break and a chlorine radical to split off from the CFC molecule. For example:



This chlorine radical then reacts with the ozone, destroying it:



One of the worrying things about this process is that, although the above equation indicates that the chlorine radical has also been destroyed, it has been found that it is then regenerated by a further reaction. It has consequently been found that a single chlorine radical is capable of destroying hundreds of thousands of ozone molecules.

This is not the only problem with CFCs. They are also significant 'greenhouse gases' (see Chapter 13, p. 212).

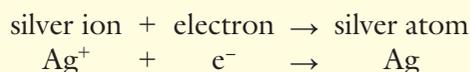
Ozone depletion and greenhouse effects have become such serious problems that an international agreement, known as the Montreal Protocol on Substances that Deplete the Ozone Layer, was agreed in 1987.

Research is now going ahead, with some success, to produce safer alternatives to CFCs. At present, better alternatives, called hydrochlorofluorocarbons (HCFCs), have been developed. These have lower ozone depletion effects and are not effective greenhouse gases. It is believed that if the

agreements are adhered to then the ozone layer will recover by 2050.

Photosynthesis is a very important reaction (Chapter 11, p. 172) which occurs only when sunlight falls on leaves containing the green pigment chlorophyll. The chlorophyll acts as a catalyst and the rate of photosynthesis depends on the intensity of the light. Another chemical reaction that takes place only in light is that which occurs in photographic film. This is a transparent plastic strip coated with emulsion: a layer of gelatin throughout which are spread many millions of tiny crystals of silver halides, in particular, silver bromide ( $\text{AgBr}$ ). The emulsion used is similar for both black-and-white and colour film.

When light hits a silver bromide crystal, silver cations ( $\text{Ag}^+$ ) accept an electron (reduction) from the bromide ions ( $\text{Br}^-$ ), which are oxidised. Hence, silver atoms and bromine atoms are produced in the emulsion.



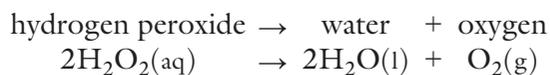
The more light that falls on the film, the greater the amount of silver deposited. Places where the most silver is deposited are darkest when the negative is developed and lightest on the photographic print.

### Question

- 1 Devise an experiment to show how sunlight affects the rate of formation of silver from the silver salts silver chloride and silver bromide.

## Catalysts

Over 90% of industrial processes use **catalysts**. A catalyst is a substance which can alter the rate of a reaction without being chemically changed itself. In the laboratory, the effect of a catalyst can be observed using the decomposition of hydrogen peroxide as an example.



The rate of decomposition at room temperature is very slow. There are substances, however, which will speed up this reaction, one being manganese(IV) oxide. When black manganese(IV) oxide powder

is added to hydrogen peroxide solution, oxygen is produced rapidly. The rate at which this occurs can be followed by measuring the volume of oxygen gas produced with time.

Some sample results from experiments of this type have been plotted in Figure 7.13. At the end of the reaction, the manganese(IV) oxide can be filtered off and used again. The reaction can proceed even faster by increasing the amount and surface area of the catalyst. This is because the activity of a catalyst involves its surface. Note that, in gaseous reactions, if dirt or impurities are present on the surface of the catalyst, it will not act as efficiently; it is said to have been 'poisoned'. Therefore, the gaseous reactants must be pure.

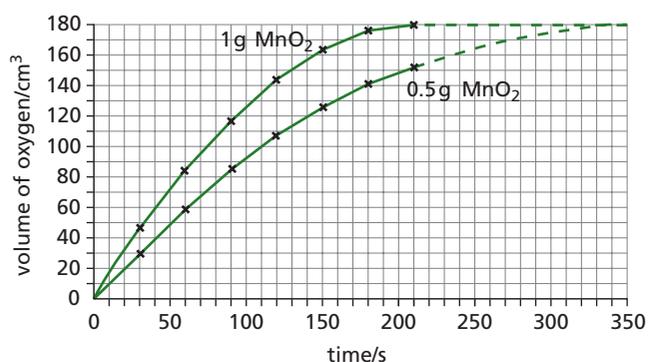


Figure 7.13 Sample data for differing amounts of MnO<sub>2</sub> catalyst.

Chemists have found that:

- a small amount of catalyst will produce a large amount of chemical change
- catalysts remain unchanged chemically after a reaction has taken place, but they can change physically. For example, a finer manganese(IV) oxide powder is left behind after the decomposition of hydrogen peroxide
- catalysts are specific to a particular chemical reaction.

Some examples of chemical processes and the catalysts used are shown in Table 7.1.

Table 7.1 Examples of catalysts.

Process	Catalyst
Haber process – for the manufacture of ammonia	Iron
Contact process – for the manufacture of sulfuric acid	Vanadium(V) oxide
Oxidation of ammonia to give nitric acid	Platinum
Hydrogenation of unsaturated oils to form fats in the manufacture of margarines	Nickel

A catalyst increases the rate of a chemical reaction by providing an alternative reaction path which has a lower activation energy. A catalyst does not increase the number of collisions between the reactant particles but only causes more of the collisions to become successful collisions, so increasing the rate of the reaction.

If the activation energy is lowered by using a catalyst then, on collision, more particles will go on to produce products at a given temperature (Figure 7.14).

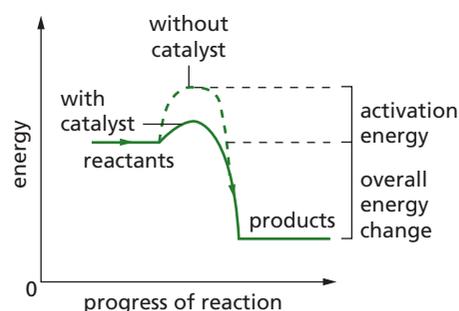
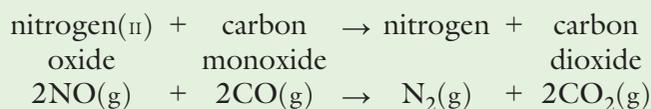
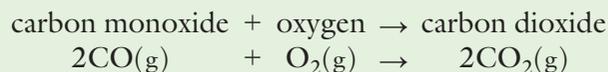


Figure 7.14 Energy level diagram showing activation energy, with and without a catalyst.

### Catalytic converters

European regulations state that all new cars have to be fitted with catalytic converters as part of their exhaust system (Figure 7.15). Car exhaust fumes contain pollutant gases such as carbon monoxide (CO) formed from the incomplete combustion of hydrocarbons in the fuel, and nitrogen(II) oxide (NO) formed by the reaction of nitrogen gas and oxygen gas from the air. The following reactions proceed of their own accord but very slowly under the conditions inside an exhaust.



The catalyst in the converter speeds up these reactions considerably. In these reactions, the pollutants are converted to carbon dioxide and nitrogen, which are naturally present in the air. The removal of oxides of nitrogen is important because they cause respiratory disease. They are also involved in the production of photochemical smogs (Figure 7.16) which occur worldwide in major cities, especially in the summer. It should be

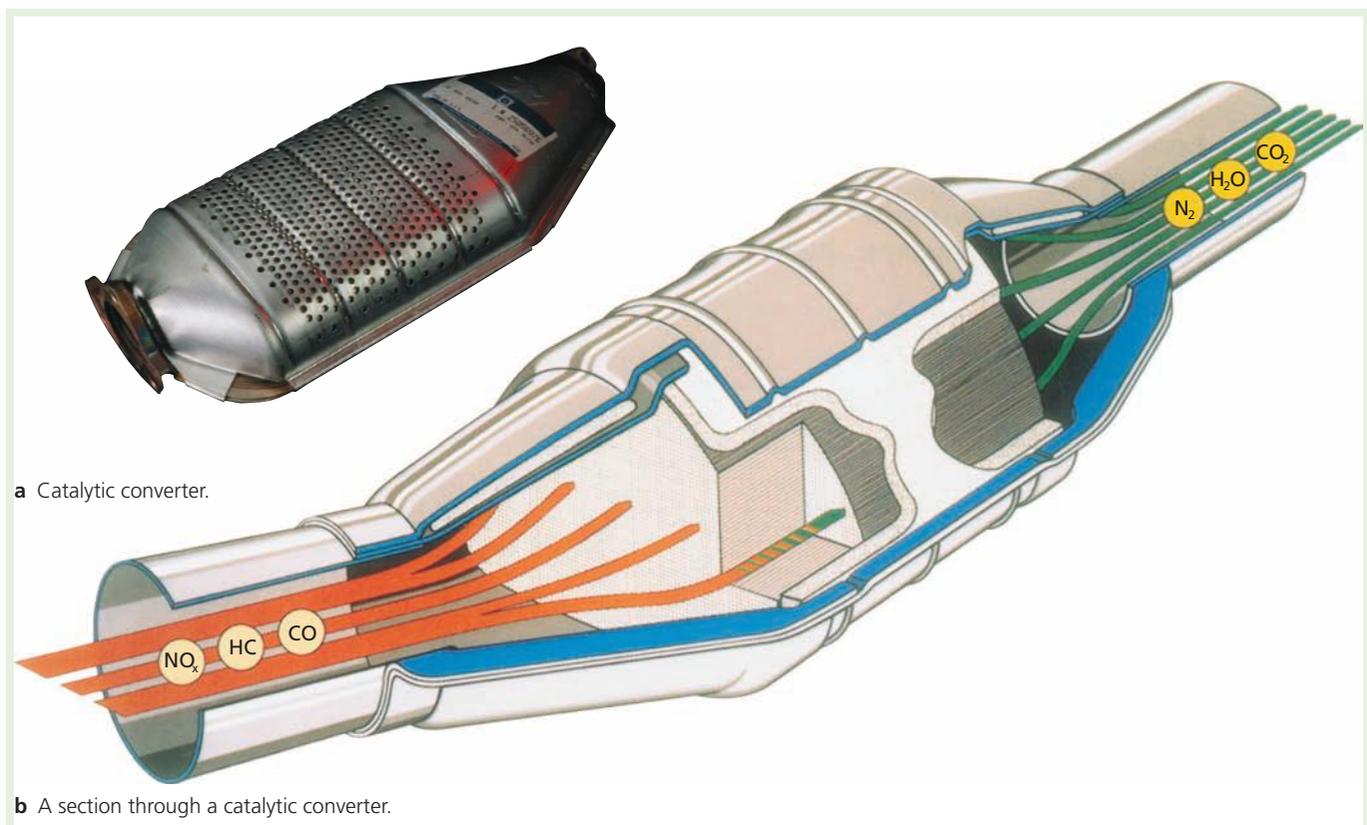


Figure 7.15



Figure 7.16 The haze is due to pollution caused mainly by cars without catalyst exhaust systems.

noted, however, that the catalytic converter can only be used with unleaded petrol and that, due to impurities being deposited on the surface of the catalyst, it becomes poisoned and has to be replaced every five or six years.

### Questions

- 1 Using a catalysed reaction of your choice, devise an experiment to follow the progress of the reaction and determine how effective the catalyst is.
- 2 Why do some people consider catalytic converters not to be as environmentally friendly as suggested in their advertising material?
- 3 Unreacted hydrocarbons such as octane,  $C_8H_{18}$  (from petrol), also form part of the exhaust gases. These gases are oxidised in the converter to carbon dioxide and water vapour. Write an equation for the oxidation of octane.

## Enzymes

**Enzymes** are protein molecules produced in living cells. They are catalysts which speed up hundreds of different chemical reactions going on inside living cells. These biological catalysts are very specific in that each chemical reaction has a different enzyme

catalyst. There are literally hundreds of different kinds of enzyme. Enzymes all have an active site. The active site is a particular shape and locks into a corresponding shape in a reactant molecule. When this has happened, the enzyme can work to break up the reactant (Figure 7.17).



**Figure 7.17** The enzyme molecules (red, pink, green and blue) lock exactly to a particular reactant molecule (yellow). Once the enzyme is locked on, the reactant molecule breaks up into pieces.

For example, hydrogen peroxide is produced within our bodies. However, it is extremely damaging and must be decomposed very rapidly. Catalase is the enzyme which converts hydrogen peroxide into harmless water and oxygen within our livers:

hydrogen peroxide  $\xrightarrow{\text{catalase}}$  water + oxygen



Although many chemical catalysts can work under various conditions of temperature and pressure as well as alkalinity or acidity, biological catalysts operate only under very particular conditions. For example, they operate over a very narrow temperature range and if the temperature becomes too high, they become inoperative. At temperatures above about 45 °C, they denature. This means that the specific shape of the active site of the enzyme molecule changes due to the breaking of bonds. This means that the reactant molecules are no longer able to fit into the active site.

A huge multimillion-pound industry has grown up around the use of enzymes to produce new materials. Biological washing powders (Figure 7.18) contain enzymes to break down stains such as sweat, blood and egg, and they do this at the relatively low temperature of 40 °C. This reduces energy costs, because the washing water does not need to be heated as much.



**Figure 7.18** These biological washing powders contain enzymes.

There were problems associated with the early biological washing powders. Some customers suffered from skin rashes, because they were allergic to the enzymes (Figure 7.19). This problem has been overcome to a certain extent by advising that extra rinsing is required. Also, many manufacturers have placed warnings on their packets, indicating that the powder contains enzymes which may cause skin rashes.



**Figure 7.19** An allergic reaction to a biological detergent.

Other industrial processes also make use of enzymes.

- In the manufacture of baby foods, enzymes called proteases are used to ‘pre-digest’ the protein part of the baby food. This is because young babies are unable to digest protein molecules.
- The enzyme isomerase is used to convert glucose syrup to fructose syrup. Fructose syrup is much sweeter than glucose syrup and can be used as a sweetener in slimming foods as less is needed.

- In the production of yoghurt, milk is initially heated to 90 °C for 15–30 minutes to kill any bacteria in the milk. After cooling to 40 °C, a starter culture of *Lactobacillus* bacteria is added and the mixture incubated at 40 °C for eight hours (Figure 7.20). The bacteria ferment the lactose in the milk to lactic acid, which causes the milk protein to become solid.
- In cheese making, milk is initially heated to kill bacteria and then cooled. A starter culture of *Streptococcus* bacteria is then added, which coagulates the milk into curds and whey (Figure 7.21). The curds are put into steel or wooden drums and pressed and allowed to dry.



Figure 7.20 Yoghurt is incubated in these tanks, and allowed to mature.



Figure 7.21 Cheese is separated into curds and whey by the addition of bacteria. The liquid, whey, is separated from the curds which are then pressed.

In industry, enzymes are used to bring about reactions at normal temperatures and pressures that would otherwise require expensive conditions and equipment. Successful processes using enzymes need to ensure that:

- the enzyme is able to function for long periods of time by optimising the environment
- the enzyme is kept in place by trapping it on the surface of an inert solid (some processes immobilise the enzymes when the process is complete)
- continuous processes occur rather than batch processes.

### Questions

- 1 When using biological washing powders what factors have to be taken into consideration?
- 2 Enzymes in yeast are used in the fermentation of glucose. Why, when the temperature is raised to 45 °C, is very little ethanol actually produced compared with the amount formed at room temperature?

### Checklist

After studying Chapter 7 you should know and understand the following terms.

- **Activation energy** The excess energy that the reactants must acquire to permit the reaction to occur.
- **Catalyst** A substance which alters the rate of a chemical reaction without itself being chemically changed.
- **Catalytic converter** A device for converting dangerous exhaust gases from cars into less harmful emissions. For example, carbon monoxide gas is converted to carbon dioxide gas.

- **Enzymes** Protein molecules produced in living cells. They act as biological catalysts and are specific to certain reactions. They operate only within narrow temperature and pH ranges.
- **Reaction rate** A measure of the change which happens during a reaction in a single unit of time. It may be affected by the following factors:
  - surface area of the reactants
  - concentration of the reactants
  - the temperature at which the reaction is carried out
  - light
  - use of a catalyst.

# Chemical reactions

## Additional questions

- 1 Explain the following statements.
- A car exhaust pipe will rust much faster if the car is in constant use.
  - Vegetables cook faster when they are chopped up.
  - Industrial processes become more economically viable if a catalyst can be found for the reactions involved.
  - In fireworks it is usual for the ingredients to be powdered.
  - Tomatoes ripen faster in a greenhouse.
  - The reaction between zinc and dilute hydrochloric acid is slower than the reaction between zinc and concentrated hydrochloric acid.

- 2 A student performed two experiments to establish how effective manganese(IV) oxide was as a catalyst for the decomposition of hydrogen peroxide. The results below were obtained by carrying out these experiments with two different quantities of manganese(IV) oxide. The volume of the gas produced was recorded against time.

Time/s	0	30	60	90	120	150	180	210
Volume for 0.3 g/cm <sup>3</sup>	0	29	55	79	98	118	133	146
Volume for 0.5 g/cm <sup>3</sup>	0	45	84	118	145	162	174	182

- Draw a diagram of the apparatus you could use to carry out these experiments.
- Plot a graph of the results.
- Is the manganese(IV) oxide acting as a catalyst in this reaction? Explain your answer.
- (i) At which stage does the reaction proceed most quickly?  
(ii) How can you tell this from your graph?  
(iii) In terms of particles, explain why the reaction is quickest at the point you have chosen in (i).
- Why does the slope of the graph become less steep as the reaction proceeds?
- What volume of gas has been produced when using 0.3 g of manganese(IV) oxide after 50 s?
- How long did it take for 60 cm<sup>3</sup> of gas to be produced when the experiment was carried out using 0.5 g of the manganese(IV) oxide?
- Write a balanced chemical equation for the decomposition of hydrogen peroxide.

- 3 a Which of the following reaction mixtures will produce hydrogen more quickly at room temperature?
- zinc granules + dilute nitric acid
  - zinc powder + dilute nitric acid
- b Give an explanation of your answer to a.
- c Suggest two other methods by which the speed of this reaction can be altered.
- 4 A flask containing dilute hydrochloric acid was placed on a digital balance. An excess of limestone chippings was added to this acid, a plug of cotton wool was placed in the neck of the flask and the initial mass was recorded. The mass of the apparatus was recorded every two minutes. At the end of the experiment the loss in mass of the apparatus was calculated and the following results were obtained.

Time/min	0	2	4	6	8	10	12	14	16
Loss in mass/g	0	2.1	3.0	3.1	3.6	3.8	4.0	4.0	4.0

- Plot the results of the experiment.
  - Which of the results would appear to be incorrect? Explain your answer.
  - Write a balanced chemical equation to represent the reaction taking place.
  - Why did the mass of the flask and its contents decrease?
  - Why was the plug of cotton wool used?
  - How does the rate of reaction change during this reaction? Explain this using particle theory.
  - How long did the reaction last?
  - How long did it take for half of the reaction to occur?
- 5 a What is a catalyst?
- List the properties of catalysts.
  - Name the catalyst used in the following processes:
    - the Contact process
    - the Haber process
    - the hydrogenation of unsaturated fats.
  - Which series of metallic elements in the Periodic Table (p. 136) do the catalysts you have named in c belong to?
  - What are the conditions used in the industrial processes named in c? The following references will help you: Chapters 11, 12 and 14.

6 This question concerns the reaction of copper(II) carbonate with dilute hydrochloric acid. The equation for the reaction is:



- Sketch a graph to show the rate of production of carbon dioxide when an excess of dilute hydrochloric acid is added. The reaction lasts 40 s and produces 60 cm<sup>3</sup> of gas.
- Find on your graph the part which shows:
  - where the reaction is at its fastest
  - when the reaction has stopped.
- Calculate the mass of copper(II) carbonate used to produce 60 cm<sup>3</sup> of carbon dioxide. (*A<sub>r</sub>*: C = 12; O = 16; Cu = 63.5. One mole of a gas occupies 24 dm<sup>3</sup> at room temperature and pressure (rtp).)
- Sketch a further graph using the same axes to show what happens to the rate at which the gas is produced if:
  - the concentration of the acid is decreased
  - the temperature is increased.

7 European regulations state that all new cars have to be fitted with catalytic converters as part of their exhaust system.

- Why are these regulations necessary?
- Which gases are removed by catalytic converters?
- Which metals are often used as catalysts in catalytic converters?
- What does the term 'poisoned' mean with respect to catalysts?
- The latest converters will also remove unburnt petrol. An equation for this type of reaction is:



- Calculate the mass of carbon dioxide produced by 1.96 g of unburnt fuel.
  - Convert this mass of carbon dioxide into a volume measured at rtp.
  - If the average car produces 7.84 g of unburnt fuel a day, calculate the volume of carbon dioxide produced by the catalytic converter measured at rtp. (*A<sub>r</sub>*: H = 1; C = 12; O = 16. One mole of any gas occupies 24 dm<sup>3</sup> at rtp.)
- 8 Suggest practical methods by which the rate of reaction can be investigated in each of the following cases:
- magnesium reacting with hydrochloric acid
  - nitrogen monoxide reacting with oxygen.

## 8

# Acids, bases and salts

## Acids and alkalis

Theories of acids and bases  
The relative strengths of acids and bases  
Neutralising an acid

## Formation of salts

Methods of preparing soluble salts  
Methods of preparing insoluble salts  
More about salts  
Testing for different salts

## Crystal hydrates

Calculation of water of crystallisation

## Solubility of salts in water

## Titration

## Checklist

## Additional questions

## ● Acids and alkalis

All the substances shown in Figure 8.1 contain an **acid** of one sort or another. Acids are certainly all around us. What properties do these substances have which make you think that they are acids or contain acids?



**Figure 8.1** What do all these foods have in common?

The word acid means ‘sour’ and all acids possess this property. They are also:

- soluble in water
- corrosive.

**Alkalis** are very different from acids. They are the chemical ‘opposite’ of acids.

- They will remove the sharp taste from an acid.
- They have a soapy feel.

Some common alkaline substances are shown in Figure 8.2.



**Figure 8.2** Some common alkaline substances.

It would be too dangerous to taste a liquid to find out if it was acidic. Chemists use substances called **indicators** which change colour when they are added to acids or alkalis. Many indicators are dyes which have been extracted from natural sources, for example litmus.

Methyl orange, a common indicator used in titrations (see p. 129) is pink in an acid solution but changes to show a yellow colour in an alkaline solution. Some other indicators are shown in Table 8.1, along with the colours they turn in acids and alkalis.

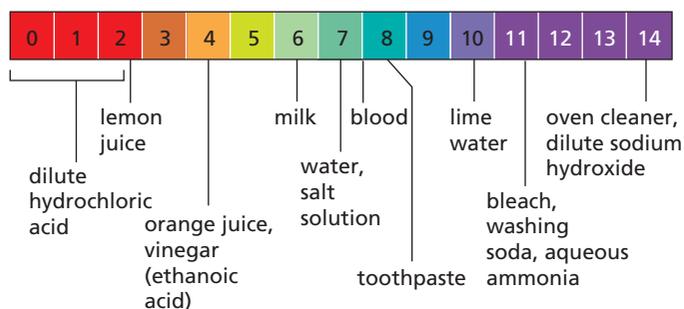
**Table 8.1** Indicators and their colours in acid and alkaline solution.

Indicator	Colour in acid solution	Colour in alkaline solution
Blue litmus	Red	Blue
Methyl orange	Pink	Yellow
Methyl red	Red	Yellow
Phenolphthalein	Colourless	Pink
Red litmus	Red	Blue

These indicators tell chemists whether a substance is acid or alkaline (Figure 8.3). To obtain an idea of how acidic or alkaline a substance is, we use another indicator known as a **universal indicator**. This indicator is a mixture of many other indicators. The colour shown by this indicator can be matched against a **pH scale**. The pH scale was developed by a Scandinavian chemist called Søren Sørensen. The pH scale runs from below 0 to 14. A substance with a pH of less than 7 is an acid. One with a pH of greater than 7 is alkaline. One with a pH of 7 is said to be neither acid nor alkaline, that is neutral. Water is the most common example of a neutral substance. Figure 8.4 shows the universal indicator colour range along with everyday substances with their particular pH values.



Figure 8.3 Indicators tell you if a substance is acid or alkaline.



a The pH scale.



b Universal indicator in solution, showing the colour range.

Figure 8.4

Another way in which the pH of a substance can be measured is by using a pH meter (Figure 8.5). The pH electrode is placed into the solution and a pH reading is given on the digital display.

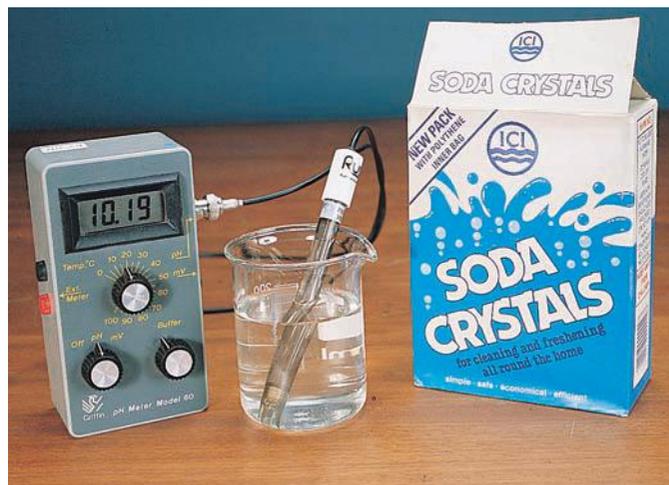


Figure 8.5 A digital pH meter.

## Theories of acids and bases

There have been many attempts to define the difference between acids and bases. The first real attempt took place in 1777, when the Frenchman Antoine Lavoisier (Figure 8.6) suggested that acids were substances that contained oxygen.

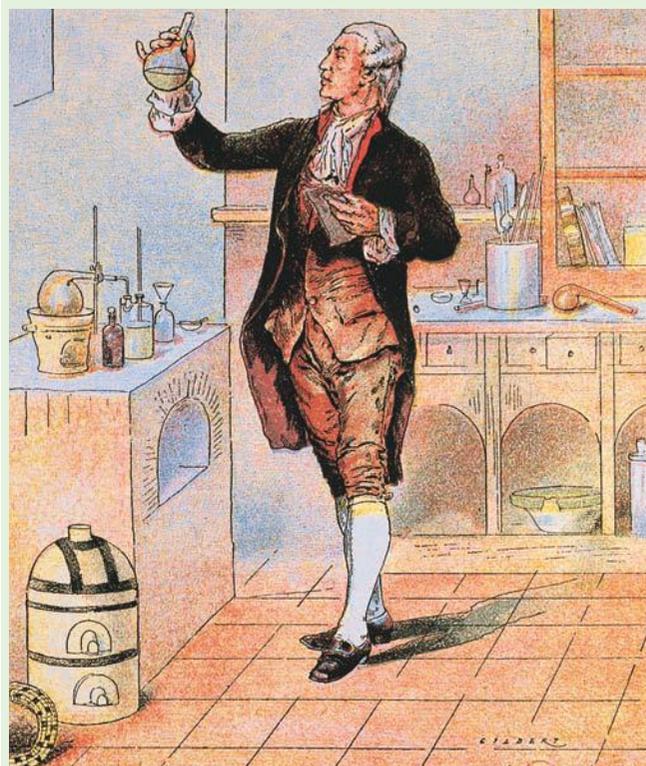


Figure 8.6 Antoine Lavoisier (1714–1794) at work in his laboratory.

It was not long after this that the ‘hydro-halic’ acids (HCl, HBr, and so on) were discovered and they had no oxygen present in them. This produced a modified theory in 1810 when the English chemist Sir Humphry Davy (1778–1829) suggested that all acids contain hydrogen as the important element; however, it was pointed out that there were many hydrogen-containing substances that were not acids. The German chemist Justus von Liebig (Figure 8.7) then made the next useful proposal about acids (1838) when he suggested that acids were substances that can react with metals to produce hydrogen gas.

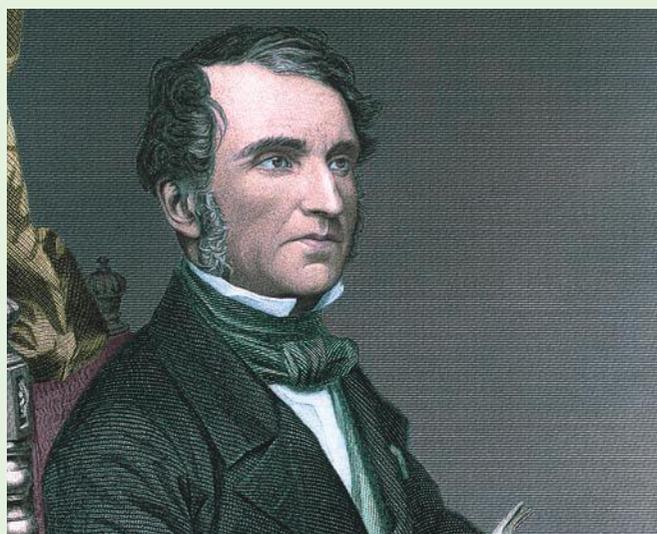


Figure 8.7 Justus von Liebig (1803–1873), who defined acids more closely.

This theory was followed in 1884 by the first really comprehensive theory of acids and bases, produced by the Swedish chemist Svante Arrhenius (1859–1927). He suggested that since these acid solutions were electrolytes (see Chapter 5) their solutions contained many ions. According to Arrhenius’ theory, acids produce hydrogen ions ( $\text{H}^+$ ) when they dissolve in water, whereas bases produce hydroxide ions ( $\text{OH}^-$ ).

It was thus recognised that water plays an important part in the acidity of a substance. This led to the suggestion that the hydrogen ion cannot exist alone in aqueous solution. This was prompted by the fact that gaseous hydrogen chloride,  $\text{HCl}(\text{g})$ , is not acidic but when it dissolves in water an acidic solution is produced.

These ideas were rather limiting since they only applied to aqueous solutions. There were situations where acid–base reactions were taking place in solvents other than water, or even in no solvent at all. This problem was addressed in 1923 by the Danish chemist Johannes Brønsted (1879–1947) and the English chemist Thomas Lowry (1874–1936) when they independently proposed a more general definition of acids and bases, and the study of acids and bases took a great step forward. This theory became known as the **Brønsted–Lowry** theory of acids and bases.

### The Brønsted–Lowry theory

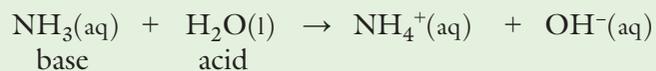
This theory defined:

- an acid as an  $\text{H}^+$  ion (or proton) donor
- a base as an  $\text{H}^+$  ion (or proton) acceptor.

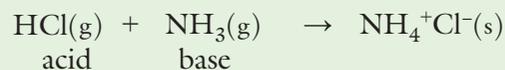
The theory explains why a pure acid behaves differently from its aqueous solution, since for an acid to behave as an  $\text{H}^+$  ion donor it must have another substance present to accept the  $\text{H}^+$  ion. So the water, in the aqueous acid solution, is behaving as a Brønsted–Lowry base and accepting an  $\text{H}^+$  ion. Generally:



If a substance can behave both as a Brønsted–Lowry acid and as a Brønsted–Lowry base then it is called **amphoteric**. Water has this ability. As well as reacting with acids (above) it can also react with Brønsted–Lowry bases such as ammonia in the following way to form the base  $\text{OH}^-$ :



The reaction between hydrogen chloride gas and ammonia can be described as an acid–base reaction under this theory. The hydrogen chloride molecule acts as a proton donor and the ammonia molecule acts as the proton acceptor (Figure 8.8).





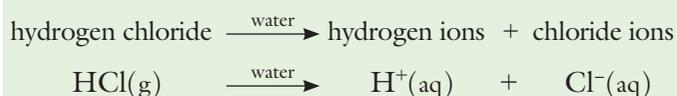
**Figure 8.8** The hydrogen chloride molecule (from concentrated hydrochloric acid) acts as a hydrogen ion donor. The ammonia molecule (from concentrated ammonia) acts as a hydrogen ion acceptor.

## The relative strengths of acids and bases

The relative strength of an acid is found by comparing one acid with another. The strength of any acid depends upon how many molecules dissociate (or ionise) when the acid is dissolved in water. The relative strength of a base is found by comparing one base with another and is again dependent upon the dissociation of the base in aqueous solution.

### Strong and weak acids

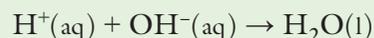
A typical strong acid is hydrochloric acid. It is formed by dissolving hydrogen chloride gas in water. In hydrochloric acid the ions formed separate completely.



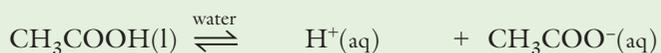
For hydrochloric acid *all* the hydrogen chloride molecules break up to form  $\text{H}^+$  ions and  $\text{Cl}^-$  ions. Any acid that behaves in this way is termed a **strong acid**. Both sulfuric acid and nitric acid also behave in this way and are therefore also termed strong acids. All these acids have a high concentration of hydrogen ions in solution ( $\text{H}^+(\text{aq})$ ) and have a low pH. Their

solutions are good conductors of electricity and they react quickly with metals, bases and metal carbonates.

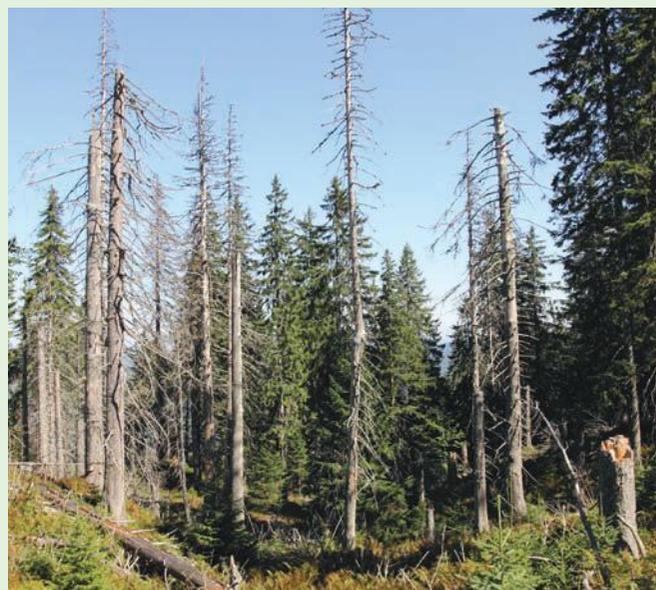
When strong acids are neutralised by strong alkalis the following reaction takes place between hydrogen ions and hydroxide ions.



A **weak acid** such as ethanoic acid, which is found in vinegar, produces few hydrogen ions when it dissolves in water compared with a strong acid of the same concentration. It is only partially ionised. Its solution has a higher pH than a strong acid, but still less than 7.



The  $\rightleftharpoons$  sign means that the reaction is **reversible**. This means that if the ethanoic acid molecule breaks down to give hydrogen ions and ethanoate ions then they will react together to re-form the ethanoic acid molecule. The fact that fewer ethanoic acid molecules dissociate compared with a strong acid, and that the reaction is reversible, means that few hydrogen ions are present in the solution. Other examples of weak acids are citric acid, found in oranges and lemons, carbonic acid, found in soft drinks, sulfurous acid (acid rain) (Figure 8.9) and ascorbic acid (vitamin C).



**Figure 8.9** Sulfurous acid is found in acid rain. It is a weak acid and is oxidised to sulfuric acid (a strong acid). Acid rain damages the environment quite badly.

Solutions of weak acids are poorer conductors of electricity and have slower reactions with metals, bases and metal carbonates.

All acids when in aqueous solution produce hydrogen ions,  $\text{H}^+(\text{aq})$ . To say an acid is a strong acid does not mean it is concentrated. The *strength* of an acid tells you how easily it dissociates (ionises) to produce hydrogen ions. The *concentration* of an acid indicates the proportions of water and acid present in aqueous solution. It is important to emphasise that a strong acid is still a strong acid even when it is in dilute solution and a weak acid is still a weak acid even if it is concentrated.

### Strong and weak bases

An alkali is a base which produces hydroxide ions,  $\text{OH}^-(\text{aq})$ , when dissolved in water. Sodium hydroxide is a **strong alkali** because when it dissolves in water its lattice breaks up completely to produce ions.

sodium hydroxide  $\xrightarrow{\text{water}}$  sodium ions + hydroxide ions



These substances which are strong alkalis produce large quantities of hydroxide ions and have a high pH. Other common strong soluble bases include potassium hydroxide.

A **weak alkali**, such as ammonia, produces fewer hydroxide ions when it dissolves in water than a strong soluble base of the same concentration. It is only partially ionised. It has a lower pH than a strong base, but still above 7.

ammonia + water  $\rightleftharpoons$  ammonium ions + hydroxide ions

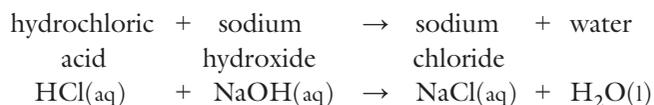


The ammonia molecules react with the water molecules to form ammonium ions and hydroxide ions. However, fewer ammonia molecules do this so only a low concentration of hydroxide ions is produced.

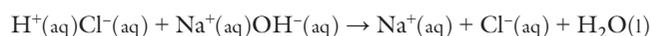
### Neutralising an acid

A common situation involving neutralisation of an acid is when you suffer from indigestion. This is caused by a build-up of acid in your stomach. Normally you treat it by taking an indigestion remedy containing a substance which will react with and neutralise the acid.

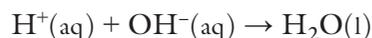
In the laboratory, if you wish to neutralise a common acid such as hydrochloric acid you can use an alkali such as sodium hydroxide. If the pH of the acid is measured as some sodium hydroxide solution is added to it, the pH increases. If equal volumes of the same concentration of hydrochloric acid and sodium hydroxide are added to one another, the resulting solution is found to have a pH of 7. The acid has been neutralised and a neutral solution has been formed.



As we have shown, when both hydrochloric acid and sodium hydroxide dissolve in water the ions separate completely. We may therefore write:



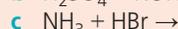
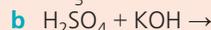
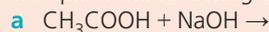
You will notice that certain ions are unchanged on either side of the equation. They are called **spectator ions** and are usually taken out of the equation. The equation now becomes:



This type of equation is known as an **ionic equation**. The reaction between any acid and alkali in aqueous solution can be summarised by this ionic equation. It shows the ion which causes acidity ( $\text{H}^+(\text{aq})$ ) reacting with the ion which causes alkalinity ( $\text{OH}^-(\text{aq})$ ) to produce neutral water ( $\text{H}_2\text{O}(\text{l})$ ).

### Questions

1 Complete the following equations:



In each case name the acid and the base. Also in parts a and b write the ionic equation for the reactions.

2 Explain the terms 'concentration' and 'strength' as applied to acids.

3 Explain what part water plays in the acidity of a solution.

4 Alongside the names of various chemicals below are shown their respective pH values in aqueous solution.

potassium hydroxide	pH 13
hydrogen bromide	pH 2
calcium hydroxide	pH 11
sodium chloride	pH 7
hydrogen chloride	pH 2
magnesium hydroxide	pH 10
citric acid	pH 4

Which of the substances is/are:

- a a strong acid?
- b a weak acid?
- c a strong alkali?
- d a weak alkali?
- e a neutral substance?

In each case write a chemical equation to show the molecules/ions present in solution.

- 5 a Write a chemical equation to represent the neutralisation of sulfuric acid by sodium hydroxide.
- b Reduce this to an ionic equation.
  - c Account for any difference you see between the ionic equation you have written and the one shown on p. 121 for the reaction of hydrochloric acid and sodium hydroxide.

## Formation of salts

In the example on p. 121, sodium chloride was produced as part of the neutralisation reaction. Compounds formed in this way are known as **normal salts**. A normal salt is a compound that has been formed when all the hydrogen ions of an acid have been replaced by metal ions or by the ammonium ion ( $\text{NH}_4^+$ ).

Normal salts can be classified as those which are soluble in water or those which are insoluble in water. The following salts are soluble in cold water:

- all nitrates
- all common sodium, potassium and ammonium salts
- all chlorides except lead, silver and mercury
- all sulfates except lead, barium and calcium.

Salts are very useful substances, as you can see from Table 8.2 and Figure 8.10.

**Table 8.2** Useful salts.

Salt	Use
Ammonium chloride	In torch batteries
Ammonium nitrate	In fertilisers
Calcium carbonate	Extraction of iron, making cement, glass making
Calcium chloride	In the extraction of sodium, drying agent (anhydrous)
Calcium sulfate	For making plaster boards, plaster casts for injured limbs
Iron(II) sulfate	In 'iron' tablets
Magnesium sulfate	In medicines
Potassium nitrate	In fertiliser and explosives manufacture
Silver bromide	In film photography
Sodium carbonate	Glass making, softening water, making modern washing powders
Sodium chloride	Making hydrochloric acid, for food flavouring, hospital saline, in the Solvay process for the manufacture of sodium carbonate
Sodium stearate	In some soaps
Tin(II) fluoride	Additive to toothpaste



a Silver bromide is used in photography.



b Iron(II) sulfate is the iron-containing substance in these iron and multivitamin tablets.

**Figure 8.10** Some uses of salts.

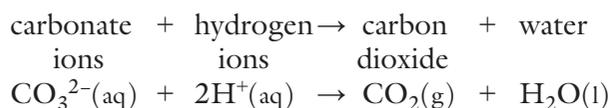
If the acid being neutralised is hydrochloric acid, salts called **chlorides** are formed. Other types of salts can be formed with other acids. A summary of the different types of salt along with the acid they have been formed from is shown in Table 8.3.

**Table 8.3** Types of salt and the acids they are formed from.

Acid	Type of salt	Example
Carbonic acid	Carbonates	Sodium carbonate ( $\text{Na}_2\text{CO}_3$ )
Ethanoic acid	Ethanoates	Sodium ethanoate ( $\text{CH}_3\text{COONa}$ )
Hydrochloric acid	Chlorides	Potassium chloride (KCl)
Nitric acid	Nitrates	Potassium nitrate ( $\text{KNO}_3$ )
Sulfuric acid	Sulfates	Sodium sulfate ( $\text{Na}_2\text{SO}_4$ )



Metal carbonates contain carbonate ions,  $\text{CO}_3^{2-}$ . In this reaction the carbonate ions react with the hydrogen ions in the acid.



### Acid + alkali (soluble base)

This method is generally used for preparing the salts of very reactive metals, such as potassium or sodium. It would certainly be too dangerous to add the metal directly to the acid. In this case, we solve the problem indirectly and use an alkali which contains the particular reactive metal whose salt we wish to prepare.

Metal oxides are **basic**. Metal oxides and hydroxides that dissolve in water to produce  $\text{OH}^-(\text{aq})$  ions are known as **alkalis**, or **soluble bases**. If the metal oxide or hydroxide does not dissolve in water it is known as an **insoluble base**.

A **base** is a substance which neutralises an acid, producing a salt and water as the only products. If the base is soluble the term alkali can be used, but there are several bases which are insoluble. It is also a substance which accepts a hydrogen ion (see p. 119). In general, most metal oxides and hydroxides (as well as ammonia solution) are bases. Some examples of soluble and insoluble bases are shown in Table 8.4. Salts can be formed by this method only if the base is soluble.

**Table 8.4** Examples of soluble and insoluble bases.

Soluble bases (alkalis)	Insoluble bases
Sodium hydroxide (NaOH)	Iron(III) oxide ( $\text{Fe}_2\text{O}_3$ )
Potassium hydroxide (KOH)	Copper(II) oxide ( $\text{CuO}$ )
Calcium hydroxide ( $\text{Ca}(\text{OH})_2$ )	Lead(II) oxide ( $\text{PbO}$ )
Ammonia solution ( $\text{NH}_3(\text{aq})$ )	Magnesium oxide ( $\text{MgO}$ )

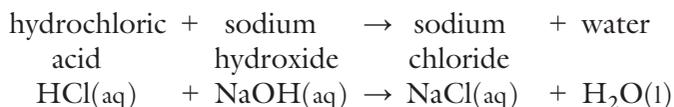
Because in this neutralisation reaction both reactants are in solution, a special technique called **titration** is required. Acid is slowly and carefully added to a measured volume of alkali using a burette (Figure 8.13) until the indicator, usually phenolphthalein, changes colour.

An indicator is used to show when the alkali has been neutralised completely by the acid. This is called the **end-point**. Once you know where the end-point is, you can add the same volume of acid to the measured volume of alkali but this time without the indicator.

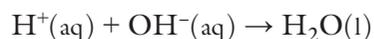


**Figure 8.13** The acid is added to the alkali until the indicator just changes colour.

The solution which is produced can then be evaporated slowly to obtain the salt. For example,

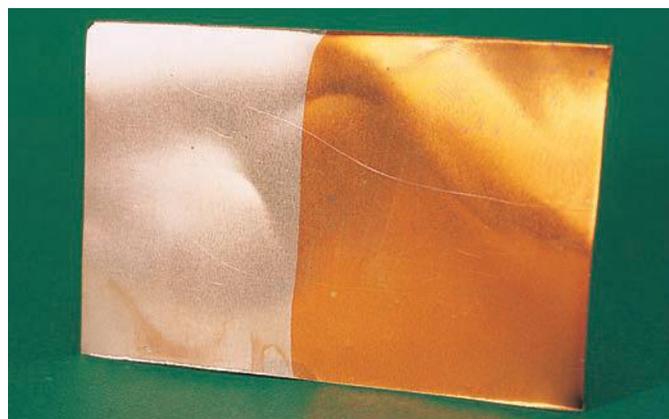


As previously discussed on p. 116, this reaction can best be described by the ionic equation:



### Acid + insoluble base

This method can be used to prepare a salt of an unreactive metal, such as lead or copper. In these cases it is not possible to use a direct reaction of the metal with an acid so the acid is neutralised using the particular metal oxide (Figure 8.14).

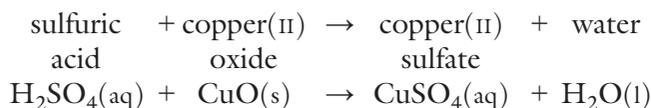


**Figure 8.14** Citric acid has been used on the right-hand side of this piece of copper metal to remove the oxide coating on its surface, giving it a shinier appearance.

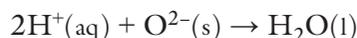
The method is generally the same as that for a metal carbonate and an acid, though some warming of the reactants may be necessary. An example of such a reaction is the neutralisation of sulfuric acid by copper(II) oxide to produce copper(II) sulfate (Figure 8.15).



**Figure 8.15** After slow evaporation to concentrate the solution, the solution is left to crystallise. Crystals of copper(II) sulfate are produced.



Metal oxides contain the oxide ion,  $\text{O}^{2-}$ . The ionic equation for this reaction is therefore:



or



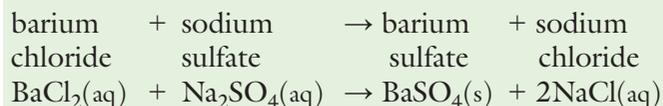
## Methods of preparing insoluble salts

The methods described above can be used to make a soluble salt, one that is soluble in water. If a salt that is insoluble in water needs to be prepared, a different technique is needed. Before we describe the technique, it is first necessary to learn which salts are soluble and which are insoluble. This can be done using the following rules of solubility:

- All nitrates are soluble in water.
- All chlorides, bromides and iodides are soluble in water, except those of lead and silver.
- All sulfates are soluble in water except barium, calcium and lead sulfates.

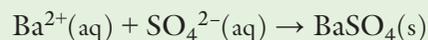
- All potassium, ammonium and sodium salts are soluble in water.
- All carbonates are insoluble, except those of potassium, ammonium and sodium.

An insoluble salt, such as barium sulfate, can be made by precipitation. In this case, solutions of the two chosen soluble salts are mixed (Figure 8.16). To produce barium sulfate, barium chloride and sodium sulfate can be used. The barium sulfate precipitate can be filtered off, washed with distilled water and dried. The reaction that has occurred is:

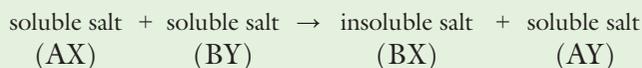


**Figure 8.16** When barium chloride solution is added to sodium sulfate a white precipitate of barium sulfate forms.

The ionic equation for this reaction is:



This method is sometimes known as **double decomposition** and may be summarised as follows:



It should be noted that even salts like barium sulfate dissolve to a very small extent. For example, 1 litre of water will dissolve  $2.2 \times 10^{-3}$  g of barium sulfate at 25 °C. This substance and substances like it are said to be **slightly soluble**.

## More about salts

You have already seen on p. 122 in Table 8.2 that salts are useful substances. Some of the salts shown in that table occur naturally and are mined, for example calcium sulfate (gypsum) and calcium carbonate (limestone). Many of the others must be made by the chemical industry, for example ammonium nitrate, iron(II) sulfate and silver bromide.

With acids such as sulfuric acid, which has two replaceable hydrogen ions per molecule, it is possible to replace only *one* of these with a metal ion. The salt produced is called an **acid salt**. An acid salt is one in which not all of the replaceable hydrogen ions of the acid have been replaced by metal ions or the ammonium ion. Some examples of acid salts are shown in Table 8.5.

**Table 8.5** Examples of acid salts.

Acid	Type of acid salt	Example
Carbonic acid (H <sub>2</sub> CO <sub>3</sub> )	Hydrogencarbonate	Sodium hydrogencarbonate (NaHCO <sub>3</sub> )
Sulfuric acid (H <sub>2</sub> SO <sub>4</sub> )	Hydrogensulfate	Potassium hydrogensulfate (KHSO <sub>4</sub> )

Sodium hydrogencarbonate is the acid salt used as the raising agent in the baking of cakes and some bread, and is often called baking soda (Figure 8.17).



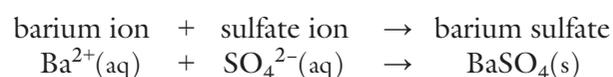
**Figure 8.17** Sodium hydrogencarbonate is used as a raising agent in bread and muffins.

## Testing for different salts

Sometimes we want to analyse a salt and find out what is in it. There are simple chemical tests which allow us to identify the anion part of the salt. These are often called **spot tests**.

### Testing for a sulfate (SO<sub>4</sub><sup>2-</sup>)

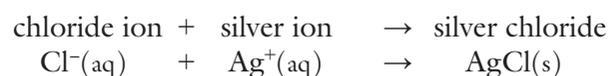
You have seen that barium sulfate is an insoluble salt (p. 125). Therefore, if you take a solution of a suspected sulfate and add it to a solution of a soluble barium salt (such as barium chloride) then a white precipitate of barium sulfate will be produced.



A few drops of dilute hydrochloric acid are also added to this mixture. If the precipitate does not dissolve, then it is barium sulfate and the unknown salt was in fact a sulfate. If the precipitate does dissolve, then the unknown salt may have been a sulfite (containing the SO<sub>3</sub><sup>2-</sup> ion).

### Testing for a chloride (Cl<sup>-</sup>), a bromide (Br<sup>-</sup>) or an iodide (I<sup>-</sup>)

Earlier in this chapter you saw that silver chloride is an insoluble salt (p. 125). Therefore, if you take a solution of a suspected chloride and add to it a small volume of dilute nitric acid, to make an **aqueous acidic solution**, followed by a small amount of a solution of a soluble silver salt (such as silver nitrate), a white precipitate of silver chloride will be produced.



If left to stand, the precipitate goes grey (Figure 8.18).



**Figure 8.18** If left to stand the white precipitate of silver chloride goes grey. This photochemical change plays an essential part in black and white photography.

In a similar way, a bromide and an iodide will react to produce either a cream precipitate of silver bromide (AgBr) or a yellow precipitate of silver iodide (AgI) (Figure 8.19).

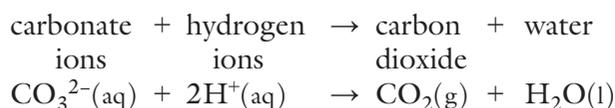


**Figure 8.19** AgCl, a white precipitate, AgBr, a cream precipitate, and AgI, a yellow precipitate.

An alternative test for iodide ions is the addition of lead nitrate solution to the iodide which results in a bright yellow precipitate of lead iodide, PbI<sub>2</sub>.

### Testing for a carbonate

If a small amount of an acid is added to some of the suspected carbonate (either solid or in solution) then effervescence occurs. If it is a carbonate then carbon dioxide gas is produced, which will turn limewater 'milky' (a cloudy white precipitate of calcium carbonate forms, see Chapter 13, p. 215).



### Testing for a nitrate

By using Devarda's alloy (45% Al, 5% Zn, 50% Cu) in alkaline solution, nitrates are reduced to ammonia. The ammonia can be identified using damp indicator paper, which turns blue.



In the reaction the nitrate ion is reduced, as oxygen is removed from the nitrogen atom, and it gains hydrogen to form ammonia, NH<sub>3</sub>. The gain of hydrogen is also a definition of reduction.

## Questions

- Complete the word equations and write balanced chemical equations for the following soluble salt preparations:
  - magnesium + sulfuric acid →
  - calcium carbonate + hydrochloric acid →
  - zinc oxide + hydrochloric acid →
  - potassium hydroxide + nitric acid →
 Also write ionic equations for each of the reactions.
- Lead carbonate and lead iodide are insoluble. Which two soluble salts could you use in the preparation of each substance? Write
  - a word equation
  - a symbol equation
  - an ionic equation
 to represent the reactions taking place.
- An analytical chemist working for an environmental health organisation has been given a sample of water which is thought to have been contaminated by a sulfate, a carbonate and a chloride.
  - Describe how she could confirm the presence of these three types of salt by simple chemical tests.
  - Write ionic equations to help you explain what is happening during the testing process.

## Crystal hydrates

Some salts, such as sodium chloride, copper carbonate and sodium nitrate, crystallise in their anhydrous forms (without water). However, many salts produce **hydrates** when they crystallise from solution. A hydrate is a salt which incorporates water into its crystal structure. This water is referred to as **water of crystallisation**. The shape of the crystal hydrate is very much dependent on the presence of water of crystallisation. Some examples of crystal hydrates are given in Table 8.6 and shown in Figure 8.20.

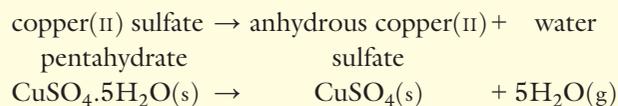
**Table 8.6** Examples of crystal hydrates.

Salt hydrate	Formula
Cobalt(II) chloride hexahydrate	CoCl <sub>2</sub> ·6H <sub>2</sub> O
Copper(II) sulfate pentahydrate	CuSO <sub>4</sub> ·5H <sub>2</sub> O
Iron(II) sulfate heptahydrate	FeSO <sub>4</sub> ·7H <sub>2</sub> O
Magnesium sulfate heptahydrate	MgSO <sub>4</sub> ·7H <sub>2</sub> O
Sodium carbonate decahydrate	Na <sub>2</sub> CO <sub>3</sub> ·10H <sub>2</sub> O
Sodium hydrogensulfate monohydrate	NaHSO <sub>4</sub> ·H <sub>2</sub> O
Sodium sulfate decahydrate	Na <sub>2</sub> SO <sub>4</sub> ·10H <sub>2</sub> O

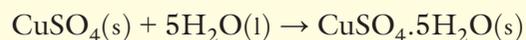


**Figure 8.20** Hydrate crystals (left to right): cobalt nitrate, calcium nitrate and nickel sulfate (top) and manganese sulfate, copper sulfate and chromium potassium sulfate (bottom).

When many hydrates are heated the water of crystallisation is driven away. For example, if crystals of copper(II) sulfate hydrate (blue) are heated strongly, they lose their water of crystallisation. Anhydrous copper(II) sulfate remains as a white powder:



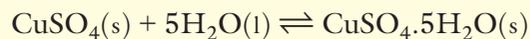
When water is added to anhydrous copper(II) sulfate the reverse process occurs. It turns blue and the pentahydrate is produced (Figure 8.21). This is an extremely exothermic process.



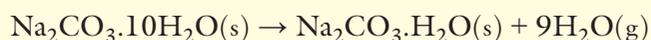
**Figure 8.21** Anhydrous copper(II) sulfate is a white powder which turns blue when water is added to it.

Because the colour change only takes place in the presence of water, the reaction is used to test for the presence of water.

These processes give a simple example of a reversible reaction:



Some crystal hydrates **effloresce**, that is they lose some or all of their water of crystallisation to the atmosphere. For example, when colourless sodium carbonate decahydrate crystals are left out in the air they become coated with a white powder, which is the monohydrate (Figure 8.22). The process is called **efflorescence**.



**Figure 8.22** A white powder forms on the surface of sodium carbonate decahydrate when it is left in the air.

With some substances, not necessarily salt hydrates, the reverse of efflorescence occurs. For example, if anhydrous calcium chloride is left in the air, it absorbs water vapour and eventually forms a very concentrated solution. This process is called **deliquescence**, and substances which behave like this are said to be **deliquescent**. Solid sodium hydroxide will deliquesce.

There are some substances which, if left out in the atmosphere, absorb moisture but do not change their state. For example, concentrated sulfuric acid, a colourless, viscous liquid, absorbs water vapour from the air and becomes a solution. Substances which do this are said to be **hygroscopic**.

## Calculation of water of crystallisation

Sometimes it is necessary to work out the percentage, by mass, of water of crystallisation in a hydrated salt. The method is the same as that used in Chapter 4, p. 65, but this time the 'H<sub>2</sub>O' is treated as an element in the calculation.

**Example**

Calculate the percentage by mass of water in the salt hydrate  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ . ( $A_r$ : H = 1; O = 16; Mg = 24; S = 32)

$$\begin{aligned} M_r \text{ for } \text{MgSO}_4 \cdot 7\text{H}_2\text{O} \\ &= 24 + 32 + (4 \times 16) + (7 \times 18) \\ &= 246 \end{aligned}$$

The mass of water as a fraction of the total mass of hydrate

$$= \frac{126}{246}$$

The percentage of water present

$$\begin{aligned} &= \frac{126}{246} \times 100 \\ &= 51.2\% \end{aligned}$$

**Questions**

- Calculate the percentage by mass of water in the following salt hydrates:
  - $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
  - $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
  - $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$ .
 ( $A_r$ : H = 1; O = 16; Na = 23; S = 32; Cu = 63.5)
- Devise an experiment to determine the percentage of water of crystallisation present in a salt hydrate of your choice.

## ● Solubility of salts in water

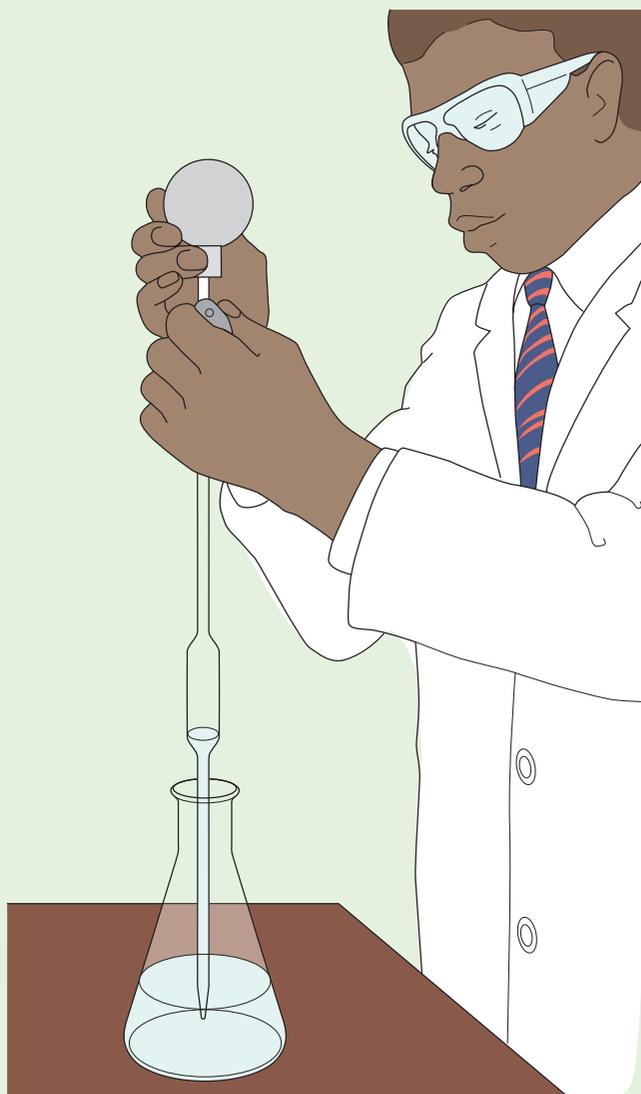
Water is a very good solvent and will dissolve a whole range of solutes, including sodium chloride and copper(II) sulfate, as well as other substances such as sugar. You can dissolve more sugar than sodium chloride in  $100 \text{ cm}^3$  of water at the same temperature. The sugar is said to be more **soluble** than the sodium chloride at the same temperature. We say that the sugar has a greater **solubility** than the sodium chloride. The solubility of a solute in water at a given temperature is the number of grams of that solute which can be dissolved in  $100 \text{ g}$  of water to produce a saturated solution at that temperature.

## ● Titration

On p. 124 you saw that it was possible to prepare a soluble salt by reacting an acid with a soluble base (alkali). The method used was that of **titration**.

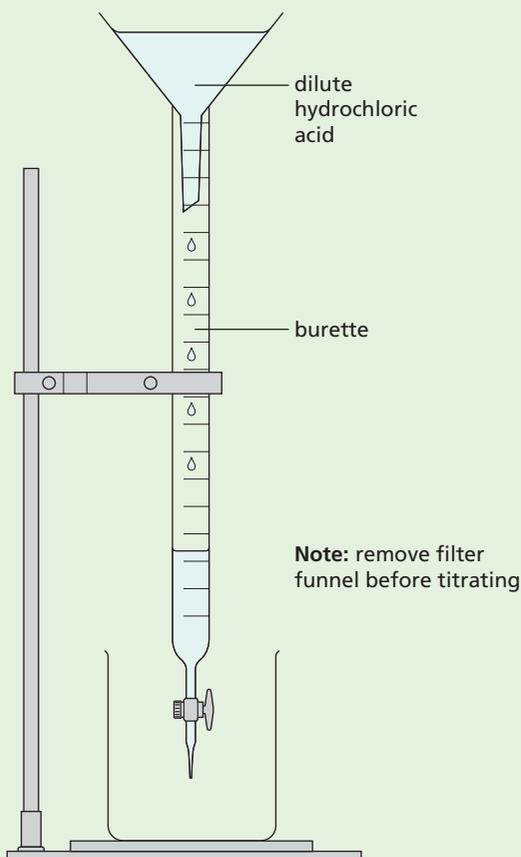
Titration can also be used to find the concentration of the alkali used. In the laboratory, the titration of hydrochloric acid with sodium hydroxide is carried out in the following way.

- $25 \text{ cm}^3$  of sodium hydroxide solution is pipetted into a conical flask to which a few drops of phenolphthalein indicator have been added (Figure 8.23). Phenolphthalein is pink in alkaline conditions but colourless in acid.



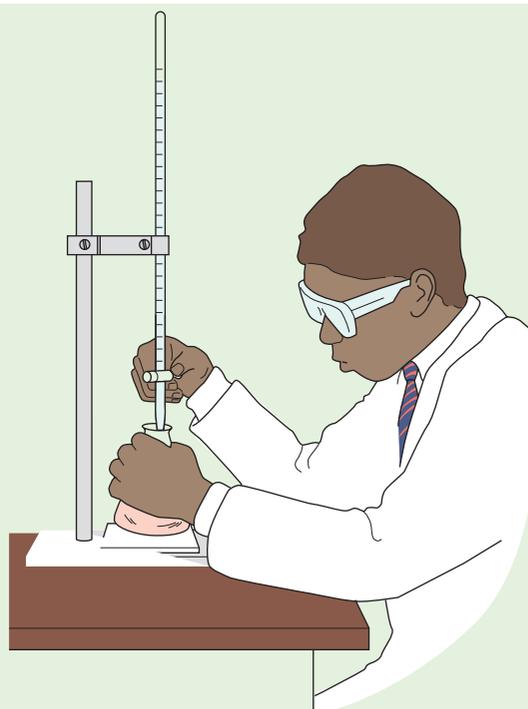
**Figure 8.23** Exactly  $25.0 \text{ cm}^3$  of sodium hydroxide solution is pipetted into a conical flask.

2 A  $0.10 \text{ mol dm}^{-3}$  solution of hydrochloric acid is placed in the burette using a filter funnel until it is filled up exactly to the zero mark (Figure 8.24).



**Figure 8.24** The burette is filled up to the zero mark with a  $0.10 \text{ mol dm}^{-3}$  solution of hydrochloric acid.

- 3 The filter funnel is now removed.
- 4 The hydrochloric acid is added to the sodium hydroxide solution in small quantities – usually no more than  $0.5 \text{ cm}^3$  at a time (Figure 8.25). The contents of the flask must be swirled after each addition of acid for thorough mixing.
- 5 The acid is added until the alkali has been neutralised completely. This is shown by the pink colour of the indicator just disappearing.



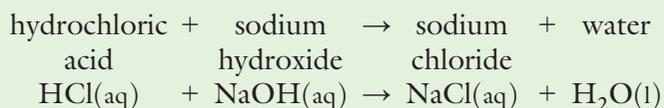
**Figure 8.25** The titration is carried out accurately.

6 The final reading on the burette at the end-point is recorded and further titrations carried out until consistent results are obtained (within  $0.1 \text{ cm}^3$  of each other). Some sample data are shown below.

Volume of sodium hydroxide solution  
=  $25.0 \text{ cm}^3$

Average volume of  $0.10 \text{ mol dm}^{-3}$  solution of hydrochloric acid added  
=  $21.0 \text{ cm}^3$

The neutralisation reaction which has taken place is:



From this equation it can be seen that 1 mole of hydrochloric acid neutralises 1 mole of sodium hydroxide.

Now you can work out the number of moles of the acid using the formula given in Chapter 4, p. 63.

$$\begin{aligned} \text{moles} &= \frac{\text{volume}}{1000} \times \text{concentration} \\ &= 21.0 \times \frac{0.10}{1000} \\ &= 2.1 \times 10^{-3} \end{aligned}$$

number of moles of hydrochloric acid = number of moles of sodium hydroxide

Therefore, the number of moles of sodium hydroxide  
 $= 2.1 \times 10^{-3}$

$2.1 \times 10^{-3}$  moles of sodium hydroxide is present in  $25.0 \text{ cm}^3$  of solution.

Therefore, in  $1 \text{ cm}^3$  of sodium hydroxide solution we have

$$\frac{2.1 \times 10^{-3}}{25.0} \text{ moles}$$

Therefore, in 1 litre of sodium hydroxide solution we have

$$\frac{2.1 \times 10^{-3}}{25.0} \times 1000 = 0.084 \text{ mole}$$

The concentration of sodium hydroxide solution is  $0.084 \text{ mol dm}^{-3}$ .

You can simplify the calculation by substituting in the following mathematical equation:

$$\frac{M_1 V_1}{M_{\text{acid}}} = \frac{M_2 V_2}{M_{\text{alkali}}}$$

where:

$M_1$  = concentration of the acid used

$V_1$  = volume of acid used ( $\text{cm}^3$ )

$M_{\text{acid}}$  = number of moles of acid shown in the chemical equation

$M_2$  = concentration of the alkali used

$V_2$  = volume of the alkali used ( $\text{cm}^3$ )

$M_{\text{alkali}}$  = number of moles of alkali shown in the chemical equation.

In the example:

$$M_1 = 0.10 \text{ mol dm}^{-3}$$

$$V_1 = 21.0 \text{ cm}^3$$

$$M_{\text{acid}} = 1 \text{ mole}$$

$$M_2 = \text{unknown}$$

$$V_2 = 25.0 \text{ cm}^3$$

$$M_{\text{alkali}} = 1 \text{ mole}$$

Substituting in the equation:

$$\frac{0.10 \times 21.0}{1} = \frac{M_2 \times 25.0}{1}$$

Rearranging:

$$M_2 = \frac{0.10 \times 21.0 \times 1}{1 \times 25.0}$$

$$M_2 = 0.084$$

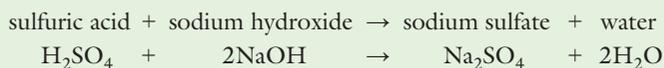
The concentration of the sodium hydroxide solution is  $0.084 \text{ mol dm}^{-3}$ .

Another example of a titration calculation could involve a neutralisation reaction in which the ratio of the number of moles of acid to alkali is not 1 : 1. The example below shows how such a calculation could be carried out.

### Example

In a titration to find the concentration of a solution of sulfuric acid,  $25 \text{ cm}^3$  of it were just neutralised by  $20.15 \text{ cm}^3$  of a  $0.2 \text{ mol dm}^{-3}$  solution of sodium hydroxide. What is the concentration of the sulfuric acid used?

First, write out the balanced chemical equation for the reaction taking place.



From this balanced equation it can be seen that 1 mole of sulfuric acid reacts with 2 moles of sodium hydroxide.

Therefore, the number of moles of sodium hydroxide used

$$= 20.15 \times \frac{0.2}{1000} = 4.03 \times 10^{-3}$$

The number of moles of sulfuric acid which will react with  $4.03 \times 10^{-3}$  moles of sodium hydroxide

$$= 4.03 \times 10^{-3} \times \frac{1}{2} = 2.015 \times 10^{-3}$$

This is the number of moles of sulfuric acid present in  $25 \text{ cm}^3$  of the solution, so the concentration of the sulfuric acid

$$= 2.015 \times 10^{-3} \times \frac{1000}{25} = 0.081 \text{ mol dm}^{-3}$$

## Questions

- 24.2 cm<sup>3</sup> of a solution containing 0.20 mol dm<sup>-3</sup> of hydrochloric acid just neutralised 25.0 cm<sup>3</sup> of a potassium hydroxide solution. What is the concentration of this potassium hydroxide solution?
- 22.4 cm<sup>3</sup> of a solution containing 0.10 mol dm<sup>-3</sup> of sulfuric acid just neutralised 25.0 cm<sup>3</sup> of a sodium hydroxide solution. What is the concentration of this sodium hydroxide solution?

## Checklist

After studying Chapter 8 you should know and understand the following terms.

- **Acid** A substance which dissolves in water, producing H<sup>+</sup>(aq) ions as the only positive ions. An acid is a hydrogen ion (H<sup>+</sup>) donor.
- **Acid salt** A substance formed when only some of the replaceable hydrogen of an acid is replaced by metal ions or the ammonium ion (NH<sub>4</sub><sup>+</sup>).
- **Alkali** A soluble base which produces OH<sup>-</sup>(aq) ions in water.
- **Base** A substance which neutralises an acid, producing a salt and water as the only products. A base is a hydrogen ion (H<sup>+</sup>) acceptor.
- **Double decomposition** The process by which an insoluble salt is prepared from solutions of two suitable soluble salts.
- **Efflorescence** The process during which a substance loses water of crystallisation to the atmosphere.
- **Hygroscopic** The ability to absorb water vapour from the atmosphere without forming solutions or changing state, for example, concentrated sulfuric acid.
- **Indicator** A substance used to show whether a substance is acidic or alkaline (basic), for example phenolphthalein.
- **Ionic equation** The simplified equation of a reaction which we can write if the chemicals involved are ionic substances.
- **Neutralisation** The process in which the acidity or alkalinity of a substance is destroyed. Destroying acidity means removing H<sup>+</sup>(aq) by reaction with a base, carbonate or metal. Destroying alkalinity means removing the OH<sup>-</sup>(aq) by reaction with an acid.
 
$$\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$$
- **Normal salt** A substance formed when all the replaceable hydrogen of an acid is completely replaced by metal ions or the ammonium ion (NH<sub>4</sub><sup>+</sup>).
- **pH scale** A scale running from 0 to 14, used for expressing the acidity or alkalinity of a solution.
- **Salt hydrates** Salts containing water of crystallisation.

- **Solubility** The solubility of a solute in a solvent at a given temperature is the number of grams of that solute which can dissolve in 100 g of solvent to produce a saturated solution at that temperature.
- **Strong acid** An acid which produces a high concentration of H<sup>+</sup>(aq) ions in water solution, for example hydrochloric acid.
- **Strong alkali** An acid which produces a high concentration of OH<sup>-</sup>(aq) ions in water solution, for example sodium hydroxide.
- **Testing for a carbonate** If effervescence occurs when an acid is added to the suspected carbonate and the gas produced tests positively for carbon dioxide, the substance is a carbonate.
- **Testing for a chloride** If a white precipitate is produced when dilute nitric acid and silver nitrate solution are added to the suspected chloride, the solution contains a chloride.
- **Testing for a sulfate** If a white precipitate is produced when dilute hydrochloric acid and barium chloride solution are added to the suspected sulfate, the solution contains a sulfate.
- **Titration** A method of volumetric analysis in which a volume of one reagent (for example an acid) is added to a known volume of another reagent (for example an alkali) slowly from a burette until an end-point is reached. If an acid and alkali are used, then an indicator is used to show that the end-point has been reached.
- **Water of crystallisation** Water incorporated into the structure of substances as they crystallise, for example in copper(II) sulfate pentahydrate (CuSO<sub>4</sub>·5H<sub>2</sub>O).
- **Weak acid** An acid which produces a low concentration of H<sup>+</sup>(aq) in water solution, for example ethanoic acid. It is only partially ionised.
- **Weak alkali** An acid which produces a low concentration of OH<sup>-</sup>(aq) in water solution, for example ammonia solution. It is only partially ionised.

# Acids, bases and salts

## Additional questions

1 Explain, with the aid of examples, what you understand by the following terms:

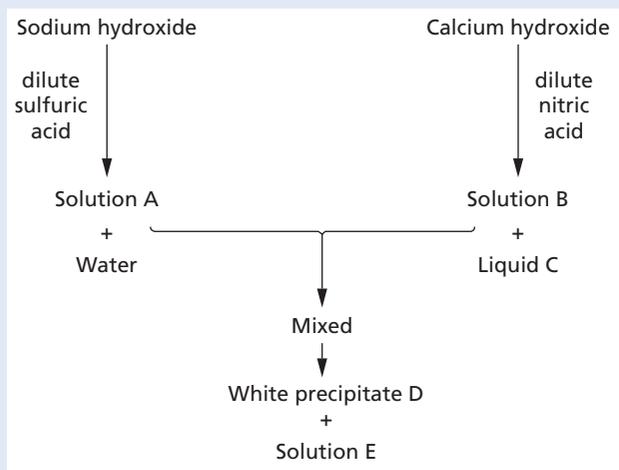
- a strong acid
- b weak acid
- c strong alkali
- d weak alkali
- e concentrated acid.

2 a Copy out and complete the table, which covers the different methods of preparing salts.

Method of preparation	Name of salt prepared	Two substances used in the preparation
Acid + alkali	Potassium sulfate	..... and .....
Acid + metal	.....	..... and dilute hydrochloric acid
Acid + insoluble base	Magnesium sulfate	..... and .....
Acid + carbonate	Copper .....	..... and .....
Precipitation	Lead iodide	..... and .....

b Write word and balanced chemical equations for each reaction shown in your table. Also write ionic equations where appropriate.

3 Study the following scheme.



- a Give the names and formulae of substances **A** to **E**.
  - b Describe a test which could be used to identify the presence of water.
  - c Which indicator is suitable for the initial reaction between the hydroxides and the dilute acids shown?
  - d Write balanced chemical equations for the reactions taking place in the scheme.
  - e Write an ionic equation for the production of the white precipitate **D**.
- 4 In a titration involving  $24.0\text{ cm}^3$  potassium hydroxide solution against a solution containing  $1\text{ mol dm}^{-3}$  of sulfuric acid,  $28.0\text{ cm}^3$  of the acid was found to just neutralise the alkali completely.
- a Write a word and balanced chemical equation for the reaction.
  - b Name a suitable indicator for the titration and state the colour change you would observe.
  - c Calculate the concentration of the alkali in  $\text{mol dm}^{-3}$ .
  - d Describe a chemical test which you could use to identify the type of salt produced during the reaction.
- 5 Explain the following, with the aid of examples:
- a neutralisation
  - b titration
  - c soluble salt
  - d insoluble salt.
- 6 Read the following passage and then answer the questions which follow.
- Sodium carbonate decahydrate *effloresces* quite readily. With some substances, such as solid sodium hydroxide, the reverse of efflorescence occurs – they *deliquesce*. There are some substances, such as concentrated sulfuric acid, which when left open to the atmosphere are diluted – they are *hygroscopic*.
- a What are the meanings of the terms in italics?
  - b What precautions should be taken to ensure that these substances are not involved in the processes described above?
  - c Which of the other salts shown in Table 8.6 on p. 127 are likely to effloresce? Give a reason for your answer.

- 7 Copper(II) sulfate crystals exist as the *pentahydrate*,  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ . It is a salt *hydrate*. If it is heated quite strongly, the *water of crystallisation* is driven off and the *anhydrous* salt remains.
- a Explain the meaning of the terms shown in italics.
- b Describe the experiment you would carry out to collect a sample of the water given off when the salt hydrate was heated strongly. Your description should include a diagram of the apparatus used and a chemical equation to represent the process taking place.
- c Describe a chemical test you could carry out to show that the colourless liquid given off was water.
- d Describe one other test you could carry out to show that the colourless liquid obtained in this experiment was **pure** water.
- e Sometimes it is necessary to work out the percentage by mass of water of crystallisation as well as the number of moles of water present in a hydrated crystal.
- (i) Use the information given to calculate the percentage, by mass, of water of crystallisation in a sample of hydrated magnesium sulfate.
- |   |           |
|---|-----------|
| Mass of crucible                            | = 14.20 g |
| Mass of crucible + hydrated $\text{MgSO}_4$ | = 16.66 g |
| Mass after heating                          | = 15.40 g |
- (ii) Calculate the number of moles of water of crystallisation driven off during the experiment as well as the number of moles of anhydrous salt remaining.  
( $A_r$ : H = 1; O = 16; Mg = 24; S = 32)
- (iii) Using the information you have obtained in (ii), write down, in the form  $\text{MgSO}_4 \cdot x\text{H}_2\text{O}$ , the formula of hydrated magnesium sulfate.

## 9

## The Periodic Table

Development of the Periodic Table

Electronic structure and the Periodic Table

Group I – the alkali metals

Group II – the alkaline earth metals

Group VII – the halogens

Displacement reactions

Group 0 – the noble gases

Transition elements

The position of hydrogen

Checklist

Additional questions

## ● Development of the Periodic Table

The Periodic Table is a vital tool used by chemists to predict the way in which elements react during chemical reactions. It is a method of categorising elements according to their properties. Scientists started to look for a way in which to categorise the known elements around 200 years ago.

The Periodic Table was devised in 1869 by the Russian Dmitri Mendeleev, who was the Professor of Chemistry at St Petersburg University (Figure 9.1). His periodic table was based on the chemical and physical properties of the 63 elements that had been discovered at that time.

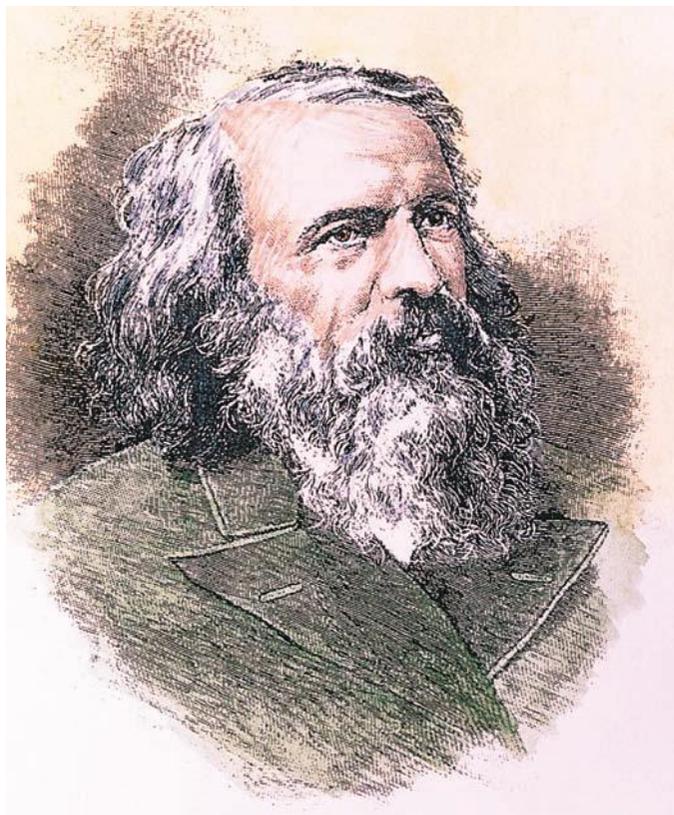


Figure 9.1 Dmitri Mendeleev (1834–1907).

However, other scientists had also attempted to categorise the known elements. In 1817, Johann Döbereiner noticed that the atomic weight (now called atomic mass) of strontium fell midway between the weights of calcium and barium. These were elements which possessed similar chemical properties. They formed a **triad** of elements. Other triads were also discovered, composed of:

chlorine, bromine, iodine

lithium, sodium, potassium

He called this the ‘**Law of Triads**’. This encouraged other scientists to search for patterns.

In 1865, John Newlands, an English chemist, arranged the 56 known elements in order of increasing atomic weight. He realised when he did this that every eighth element in the series was similar.

H Li Be B C N O F Na Mg Al Si P S Cl K

He likened this to music and called it the ‘**Law of Octaves**’. It fell down, however, because some of the weights were inaccurate and there were elements that had not been discovered then.

Mendeleev’s classification proved to be the most successful. Mendeleev arranged all the 63 known elements in order of increasing atomic weight but in such a way that elements with similar properties were in the same vertical column. He called the vertical columns **groups** and the horizontal rows **periods** (Figure 9.2). If necessary he left gaps in the table.

As a scientific idea, Mendeleev’s periodic table was tested by making predictions about elements that were unknown at that time but could possibly fill the gaps. Three of these gaps are shown by the symbols \* and † in Figure 9.2. As new elements were discovered, they were found to fit easily into the classification. For example, Mendeleev predicted the properties of the missing element ‘eka-silicon’ (†). He predicted the colour, density and melting point as well as its atomic weight.





**Figure 9.4** Transition elements have a wide range of uses, both as elements and as alloys.

The horizontal rows are called **periods** and these are numbered 1–7 going down the Periodic Table.

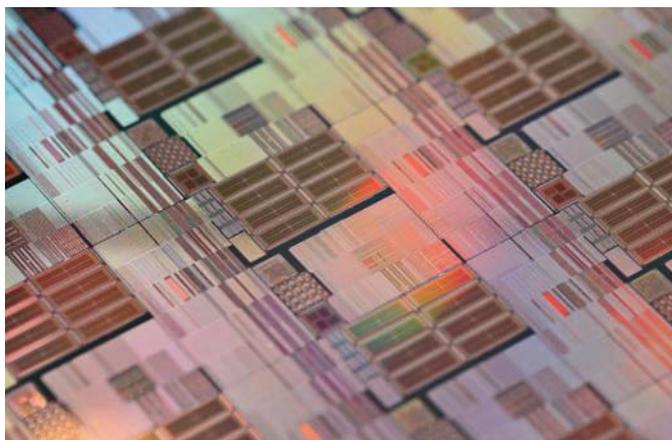
Between Groups II and III is the block of elements known as the transition elements (Figure 9.4).

The Periodic Table can be divided into two as shown by the bold line that starts beneath boron, in Figure 9.3. The elements to the left of this line are metals (fewer than three-quarters) and those on the right are non-metals (fewer than one-quarter). The elements which lie on this dividing line are known as metalloids (Figure 9.5). These elements behave in some ways as metals and in others as non-metals.

If you look at the properties of the elements across a period of the Periodic Table you will notice certain trends. For example, there is:

- a gradual change from metal to non-metal
- an increase in the number of electrons in the outer energy level of the element

- a change in the structure of the element, from giant metallic in the case of metals (e.g. magnesium, p. 56, Figure 3.43), through giant covalent (e.g. diamond, p. 52, Figure 3.33), to simple molecular (e.g. chlorine, p. 46, Figure 3.21).



**Figure 9.5** The metalloid silicon is used to make silicon ‘chips’.

## ● Electronic structure and the Periodic Table

The number of electrons in the outer energy level is discussed in Chapter 3 (p. 37). It can be seen that it corresponds with the number of the group in the Periodic Table in which the element is found. For example, the elements shown in Table 9.1 have one electron in their outer energy level and they are all found in Group I. The elements in Group 0, however, are an exception to this rule, as they have two or eight electrons in their outer energy level. The outer electrons are mainly responsible for the chemical properties of any element, and, therefore, elements in the same group have similar chemical properties (Tables 9.2 and 9.3).

**Table 9.1** Electronic structure of the first three elements of Group I.

Element	Symbol	Proton number	Electronic structure
Lithium	Li	3	2,1
Sodium	Na	11	2,8,1
Potassium	K	19	2,8,8,1

**Table 9.2** Electronic structure of the first three elements of Group II.

Element	Symbol	Proton number	Electronic structure
Beryllium	Be	4	2,2
Magnesium	Mg	12	2,8,2
Calcium	Ca	20	2,8,8,2

**Table 9.3** Electronic structure of the first three elements in Group VII.

Element	Symbol	Proton number	Electronic structure
Fluorine	F	9	2,7
Chlorine	Cl	17	2,8,7
Bromine	Br	35	2,8,18,7

The metallic character of the elements in a group increases as you move down the group. This is because electrons become easier to lose as the outer shell electrons become further from the nucleus. There is less attraction between the nucleus and the outer shell electrons because of the increased distance between them.

## ● Group I – the alkali metals

Group I consists of the five metals lithium, sodium, potassium, rubidium and caesium, and the radioactive element francium. Lithium, sodium and potassium are commonly available for use in school. They are all very reactive metals and they are stored under oil to

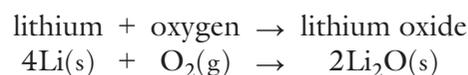
prevent them coming into contact with water or air. These three metals have the following properties.

- They are good conductors of electricity and heat.
- They are soft metals. Lithium is the hardest and potassium the softest.
- They are metals with low densities. For example, lithium has a density of  $0.53 \text{ g cm}^{-3}$  and potassium has a density of  $0.86 \text{ g cm}^{-3}$ .
- They have shiny surfaces when freshly cut with a knife (Figure 9.6).

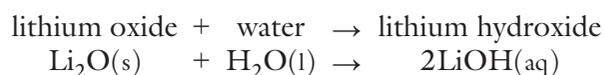


**Figure 9.6** Freshly cut sodium.

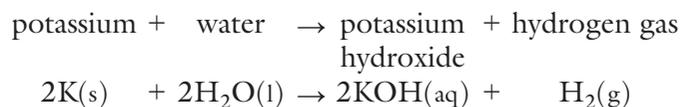
- They have low melting points. For example, lithium has a melting point of  $181^\circ\text{C}$  and potassium has a melting point of  $64^\circ\text{C}$ .
- They burn in oxygen or air, with characteristic flame colours, to form white solid oxides. For example, lithium reacts with oxygen in air to form white lithium oxide, according to the following equation:



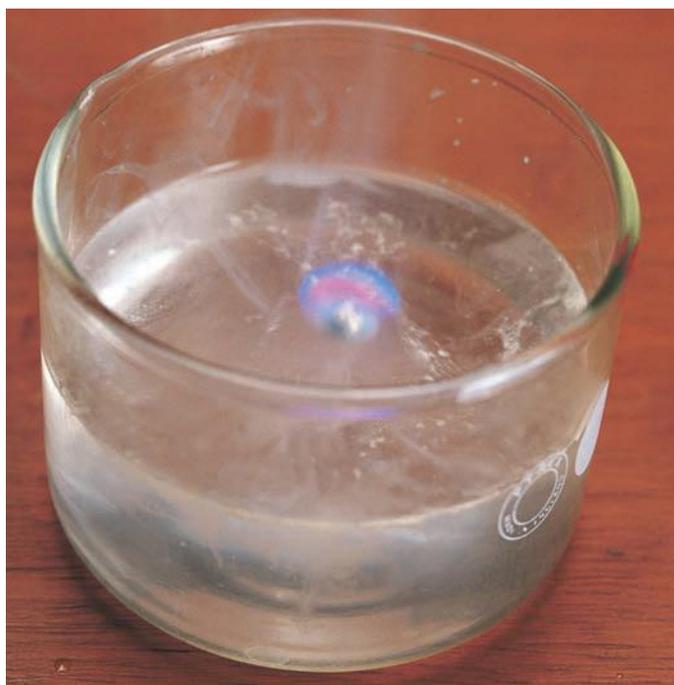
These Group I oxides all dissolve in water to form alkaline solutions of the metal hydroxide.



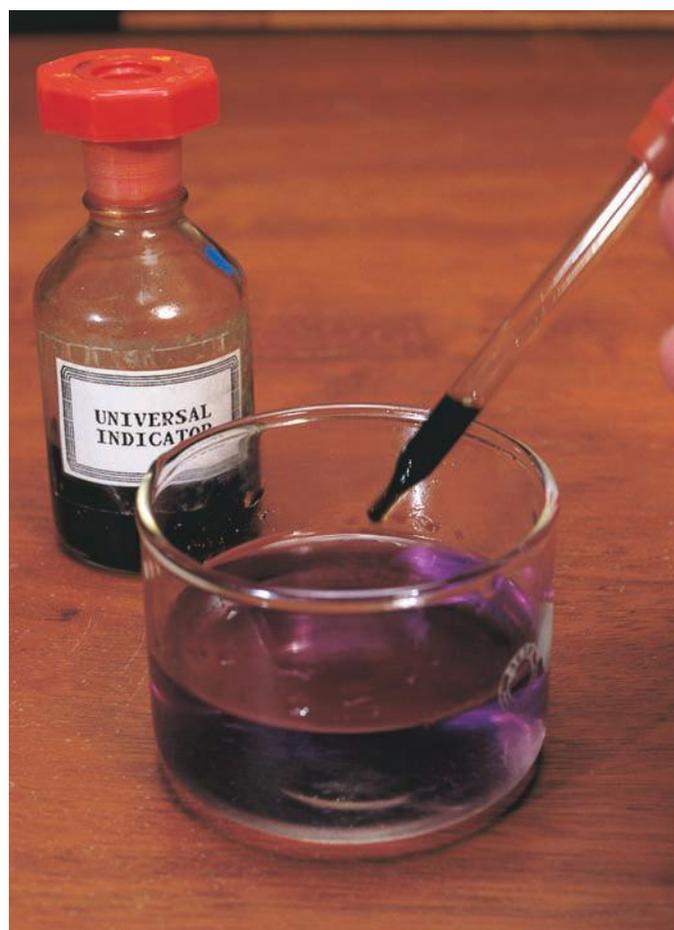
- They react vigorously with water to give an alkaline solution of the metal hydroxide as well as producing hydrogen gas. For example:



- Of these three metals, potassium is the most reactive towards water (Figure 9.7), followed by sodium and then lithium. Such gradual changes we call **trends**. Trends are useful to chemists as they



a Potassium reacts very vigorously with cold water.



b An alkaline solution is produced when potassium reacts with water.

Figure 9.7

allow predictions to be made about elements we have not observed in action.

- They react vigorously with halogens, such as chlorine, to form metal halides, for example sodium chloride (Figure 9.8).

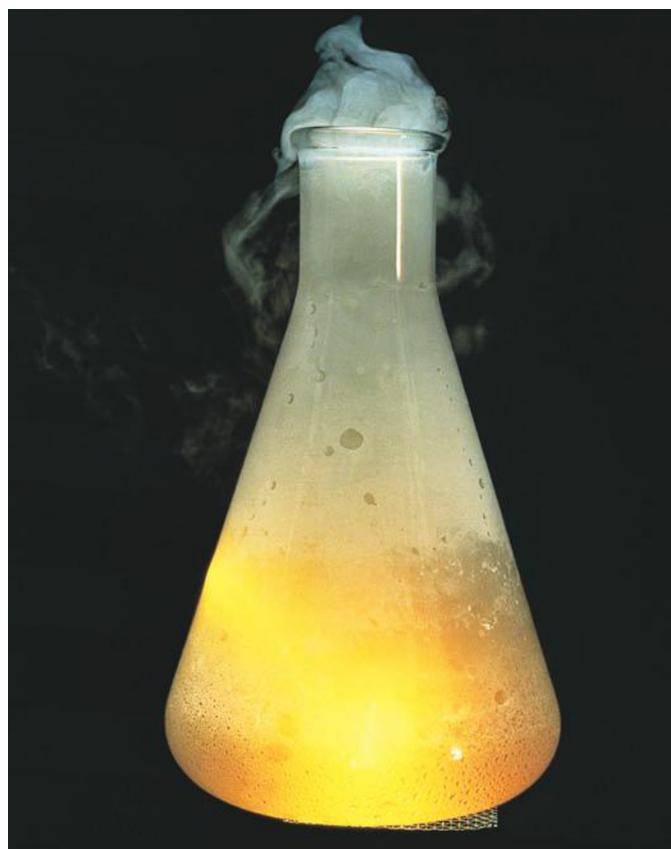
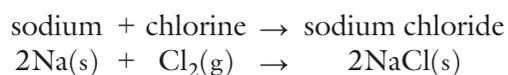


Figure 9.8 A very vigorous reaction takes place when sodium burns in chlorine gas. Sodium chloride is produced.

Considering the group as a whole, the further down the group you go the more reactive the metals become. Francium is, therefore, the most reactive Group I metal.

Table 9.1 shows the electronic structure of the first three elements of Group I. You will notice in each case that the outer energy level contains only one electron. When these elements react they lose this outer electron, and in doing so become more stable, because they obtain the electron configuration of a noble gas. You will learn more about the stable nature of these gases later in this chapter (p. 143).

When, for example, the element sodium reacts it loses its outer electron. This requires energy to overcome the electrostatic attractive forces between the outer electron and the positive nucleus (Figure 9.9).

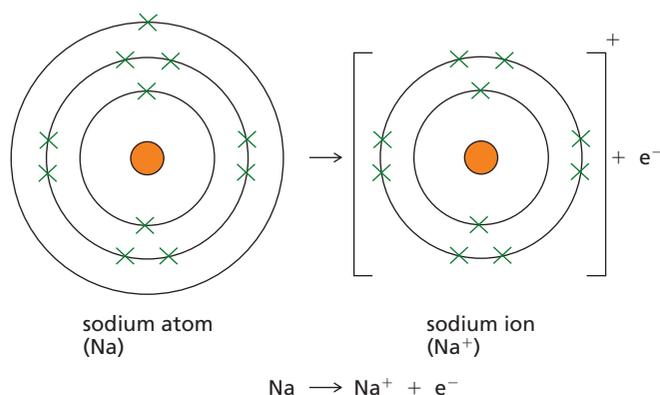


Figure 9.9 This sodium atom loses an electron to become a sodium ion.

Look at Figure 9.10. Why do you think potassium is more reactive than lithium or sodium?

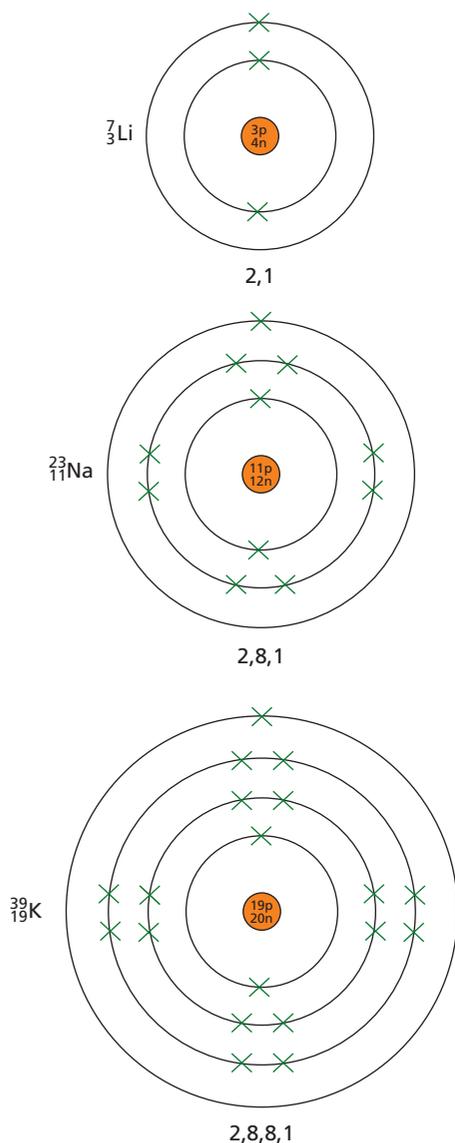


Figure 9.10 Electronic structures of lithium, sodium and potassium.

Potassium is more reactive because less energy is required to remove the outer electron from its atom than for lithium or sodium. This is because as you go down the group the size of the atoms increases and the outer electron gets further away from the nucleus and becomes easier to remove.

### Questions

- Write word and balanced chemical equations for the reactions between:
  - sodium and oxygen
  - sodium and water.
- Using the information on pp. 138–140, predict the properties of the element francium related to its melting point, density and softness. Predict how francium would react with water and write a balanced equation for the reaction.
- Write word and balanced chemical equations for the reactions between:
  - magnesium and water
  - calcium and oxygen.
- Account for the fact that calcium is more reactive than magnesium.

## Group II – the alkaline earth metals

Group II consists of the five metals beryllium, magnesium, calcium, strontium and barium, and the radioactive element radium. Magnesium and calcium are generally available for use in school. These metals have the following properties.

- They are harder than those in Group I.
- They are silvery-grey in colour when pure and clean. They tarnish quickly, however, when left in air due to the formation of a metal oxide on their surfaces (Figure 9.11).



Figure 9.11 Tarnished (left) and cleaned-up magnesium.