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Chemistry

Third Edition

Bryan Earl
Doug Wilford

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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.

Harmful
These substances are similar to toxic substances but less dangerous.

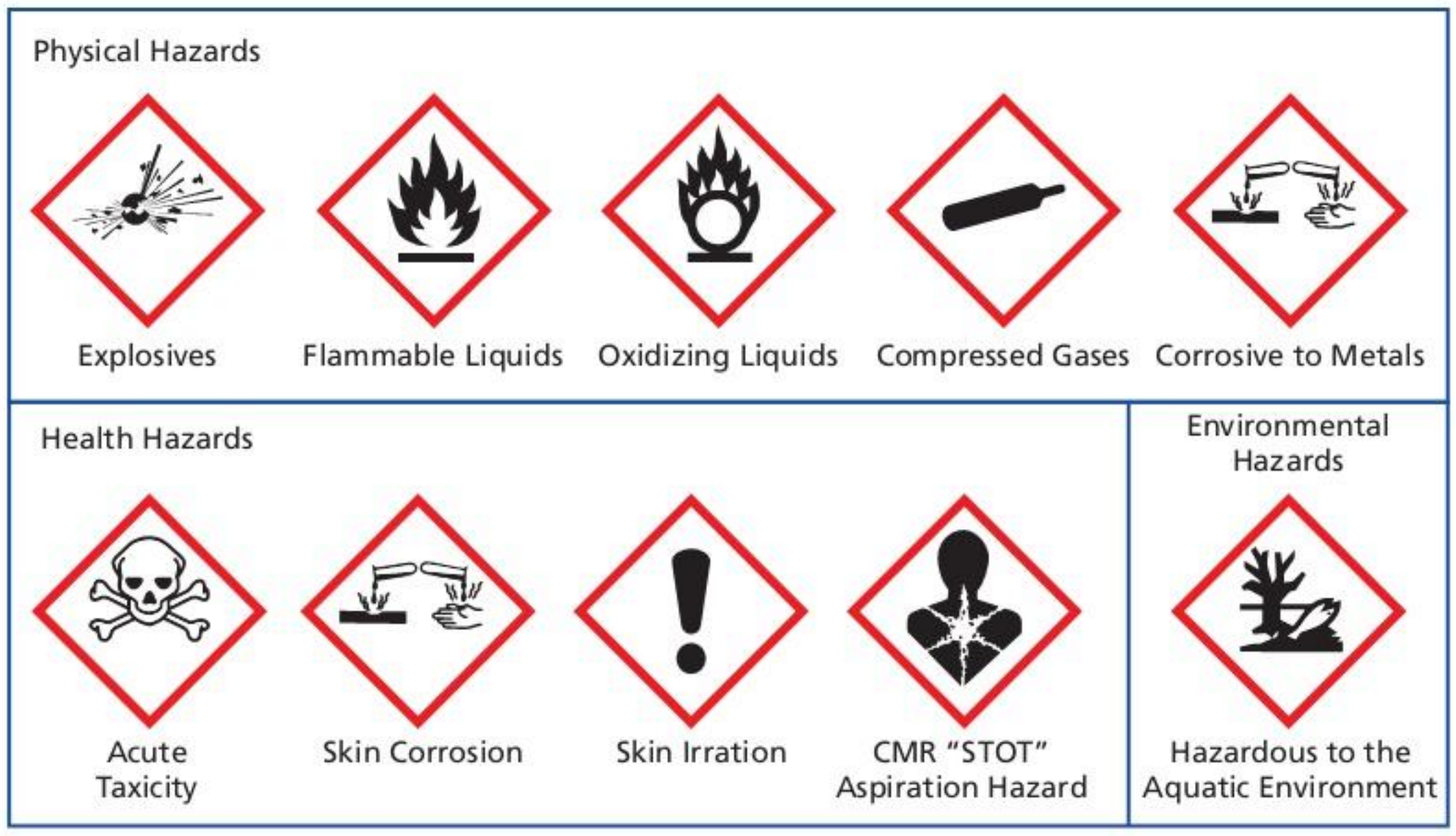
Irritant
These substances are not corrosive but can cause reddening or blistering of the skin.

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

Toxic
These substances can cause death.

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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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.

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

We use different colours to define different areas of chemistry:

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

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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.

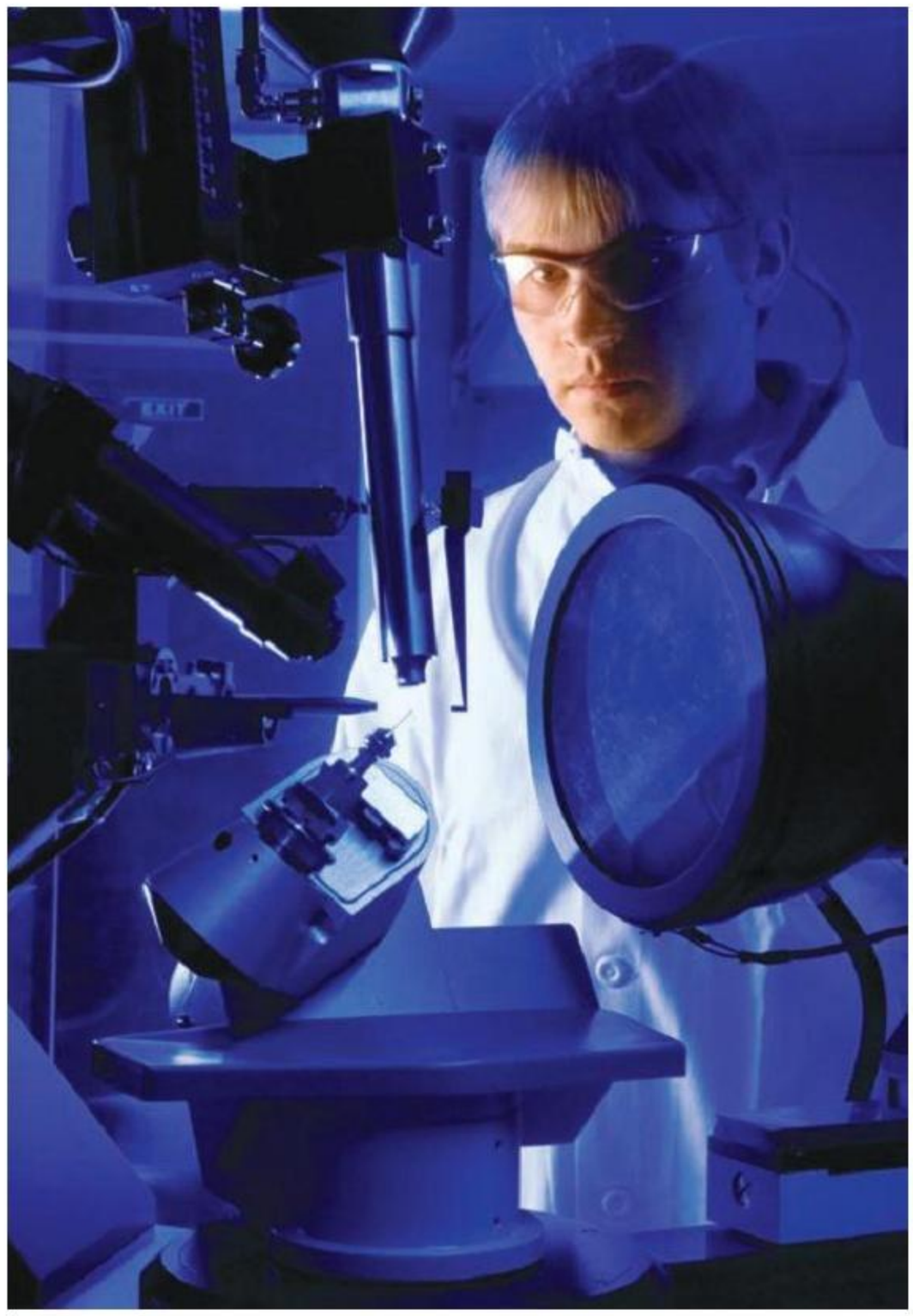


Figure 1.4 A modern X-ray crystallography instrument, used for studying crystal structure.

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.

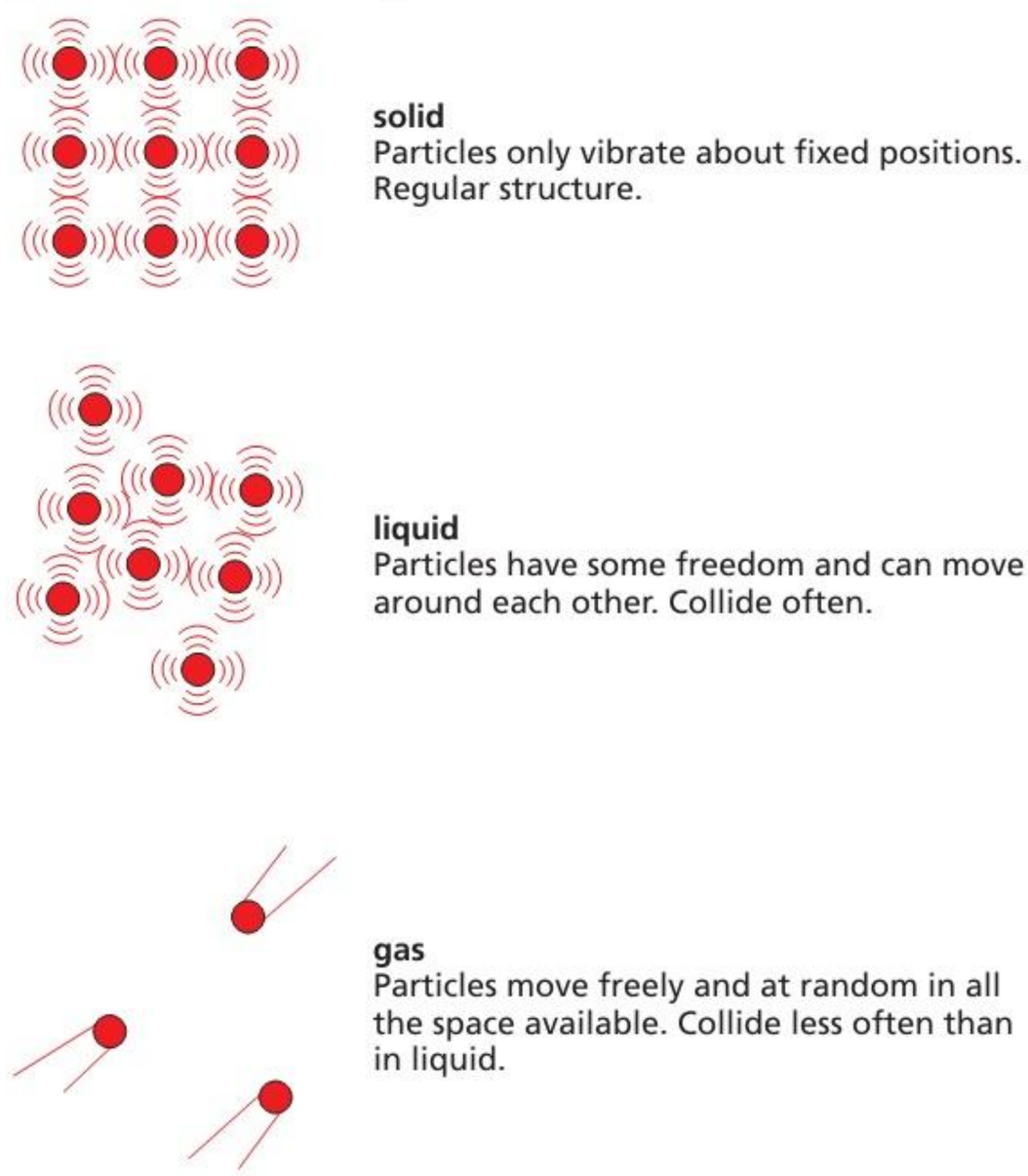


Figure 1.6 The arrangement of particles in solids, liquids and gases.

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.

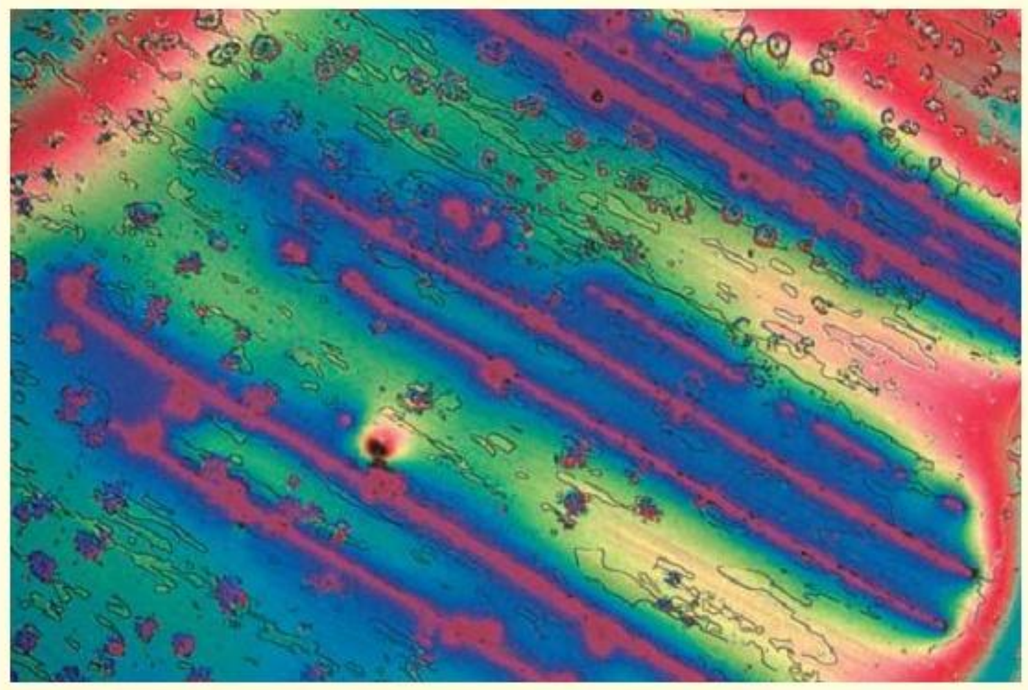


Figure 1.7 A polarised light micrograph of liquid crystals.

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.

$$\text{solid} \underset{\text{sublimation}}{\overset{\text{sublimation}}{\rightleftharpoons}} \text{gas}$$

Figure 1.9 Dry ice (solid carbon dioxide) sublimes 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 °C. When heated to just above –78 °C it changes into carbon dioxide gas. The changes of state are summarised in Figure 1.10.

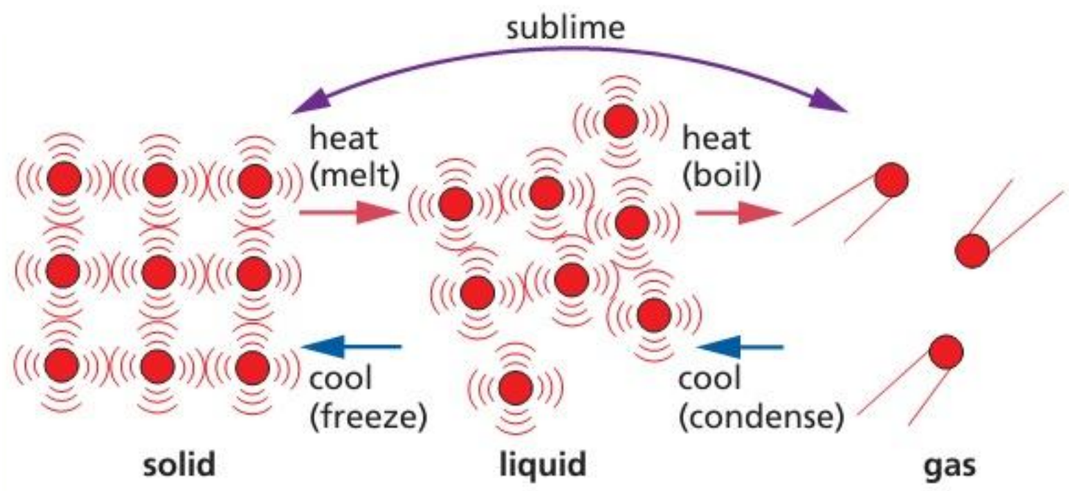


Figure 1.10 Summary of the changes of state.

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 °C to 110 °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.

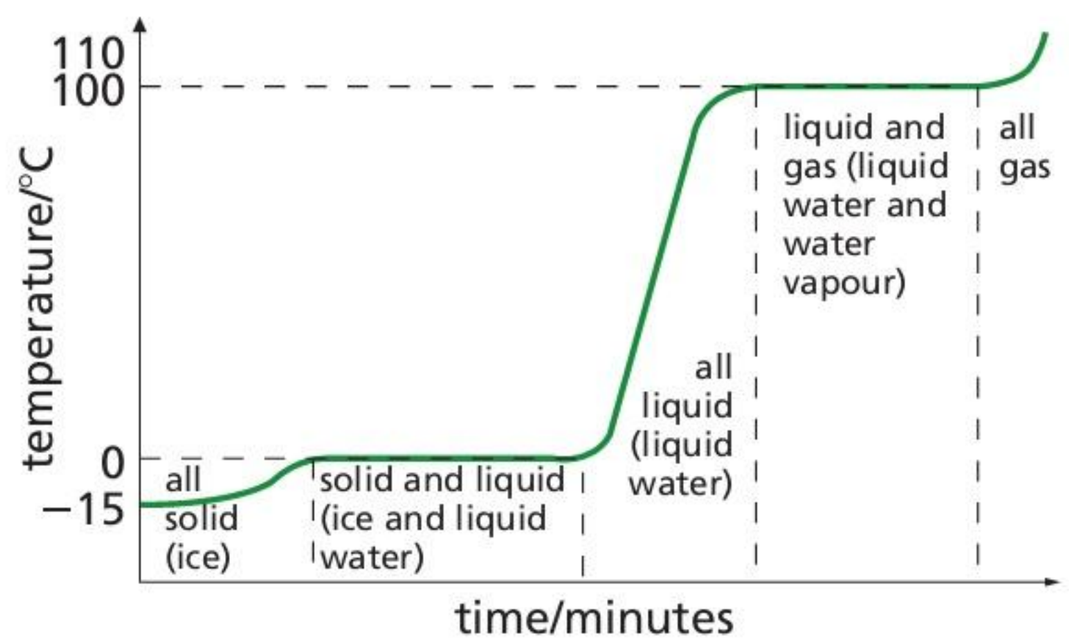


Figure 1.11 Graph of temperature against time for the change from ice at –15 °C to water to steam.

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.

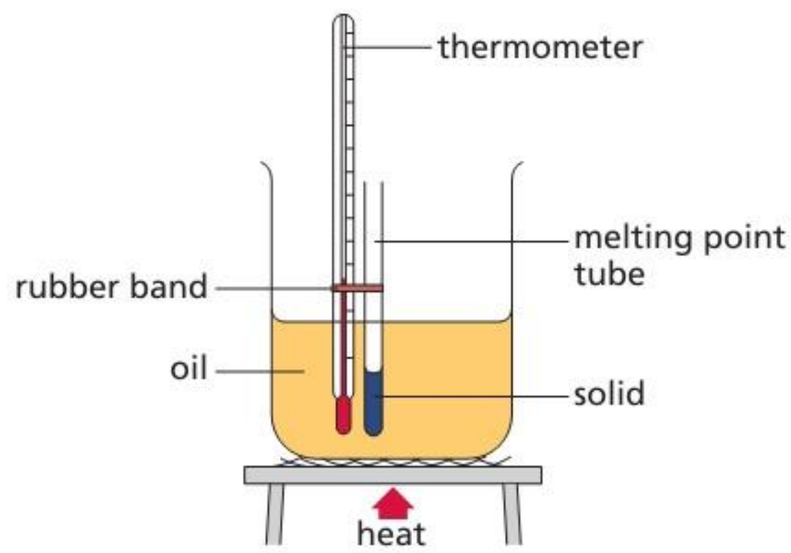


Figure 1.12 Apparatus shown here if heated slowly can be used to find the melting point of a substance such as the solid in the melting point tube.

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 °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 °C was allowed to cool to –5 °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.

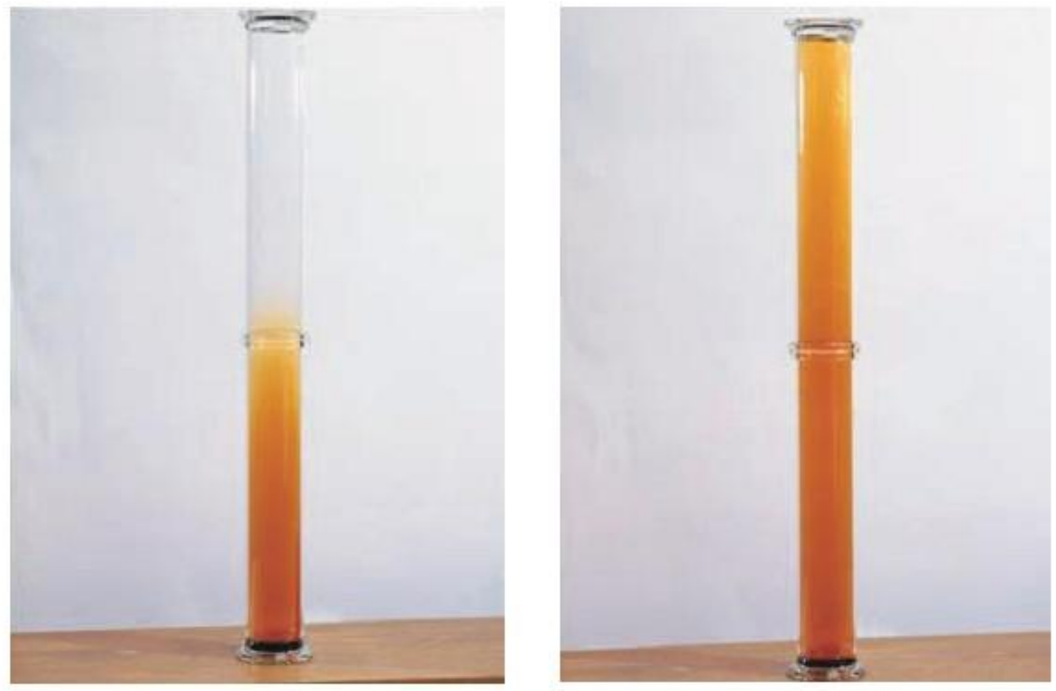


Figure 1.13 After 24 hours the bromine fumes have diffused throughout both gas jars.

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.

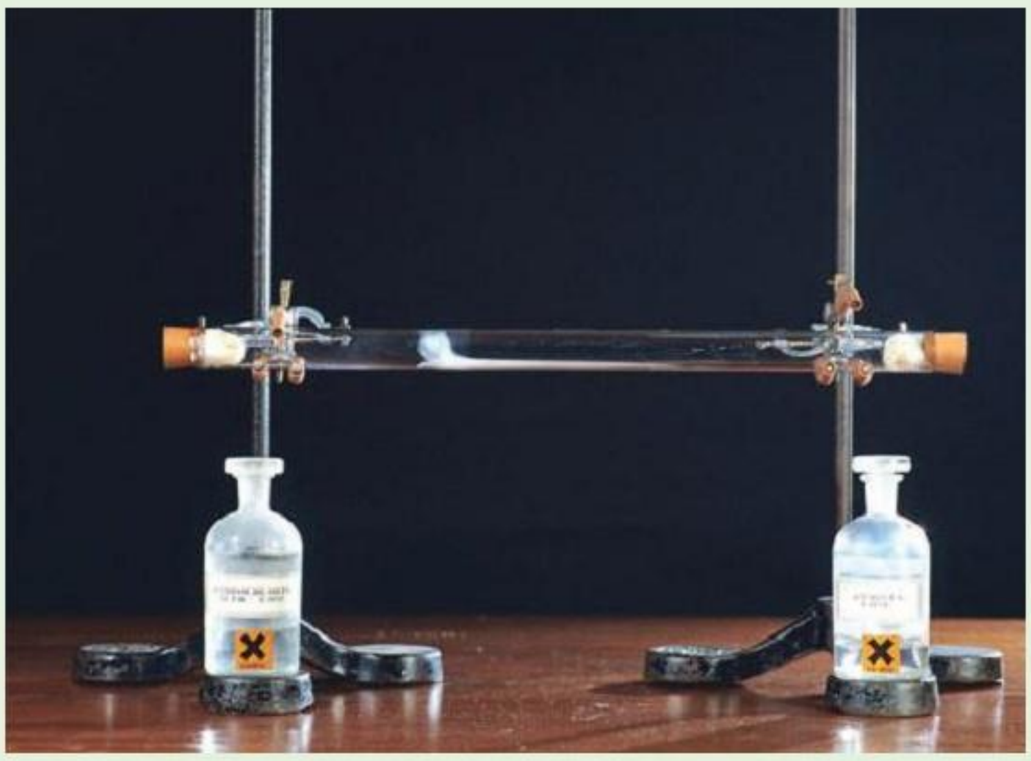


Figure 1.14 Hydrochloric acid (left) and ammonia (right) diffuse at different rates.

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).

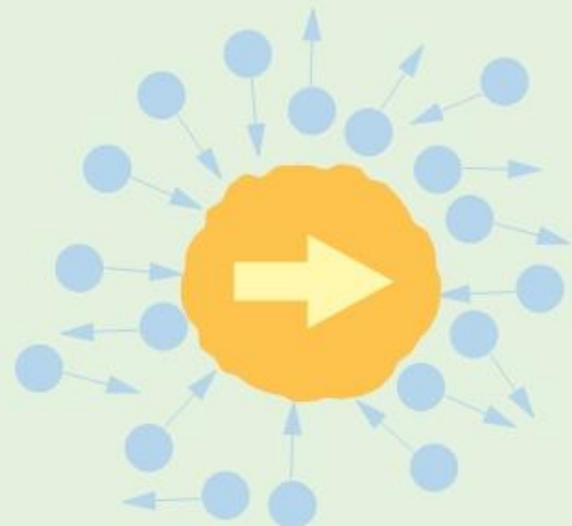


Figure 1.16a Pollen particle being bombarded by water molecules.

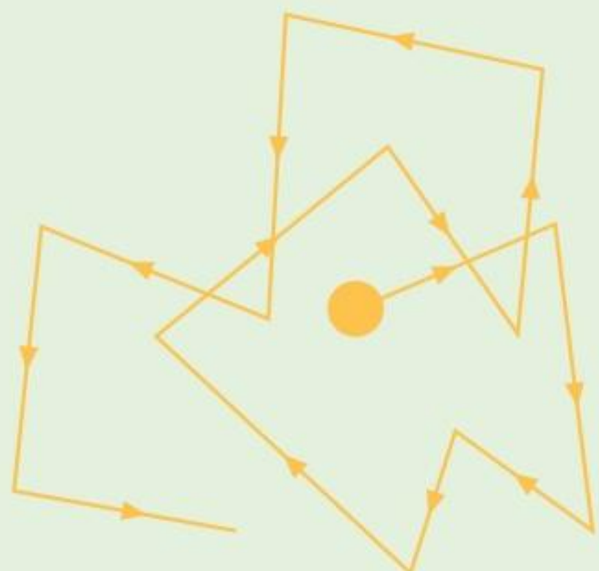


Figure 1.16b Brownian motion causes the random motion of the visible particle.

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

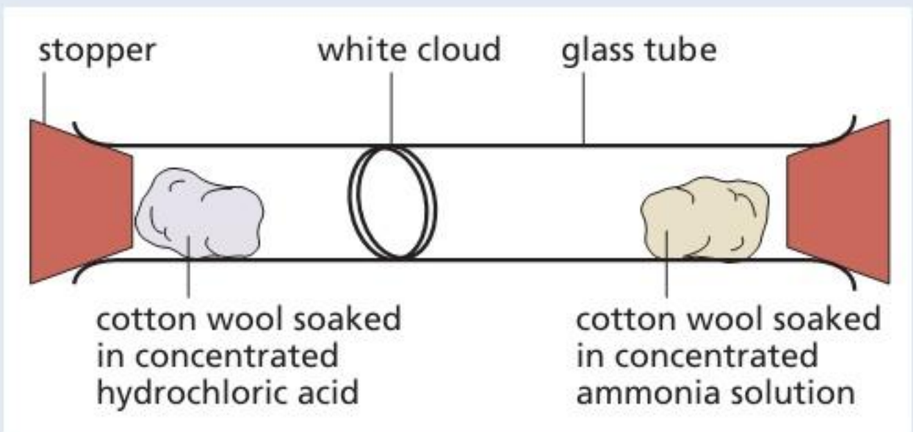
1 a Draw diagrams to show the arrangement of particles in:
(i) solid lead
(ii) molten lead
(iii) gaseous lead.
b Explain how the particles move in these three states of matter.
c Explain, using the kinetic theory, what happens to the particles in oxygen as it is cooled down.

2 Explain the meaning of each of the following terms. In your answer include an example to help with your explanation.
a Expansion.
b Contraction.
c Physical change.
d Sublimation.
e Diffusion.
f Random motion.

3 a Why do solids not diffuse?
b Give two examples of diffusion of gases and liquids found in the house.

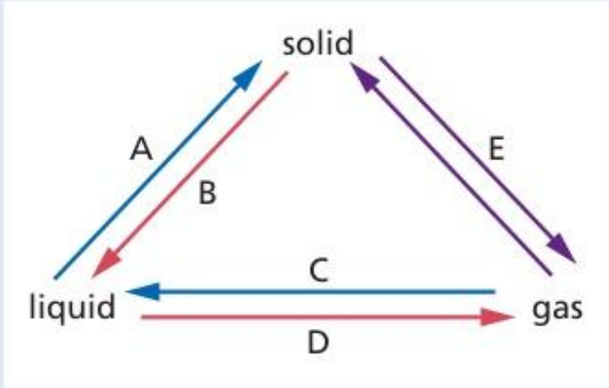
4 Use the kinetic theory to explain the following:
a When you take a block of butter out of the fridge, it is quite hard. However, after 15 minutes it is soft enough to spread.
b When you come home from school and open the door you can smell your tea being cooked.
c A football is blown up until it is 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.



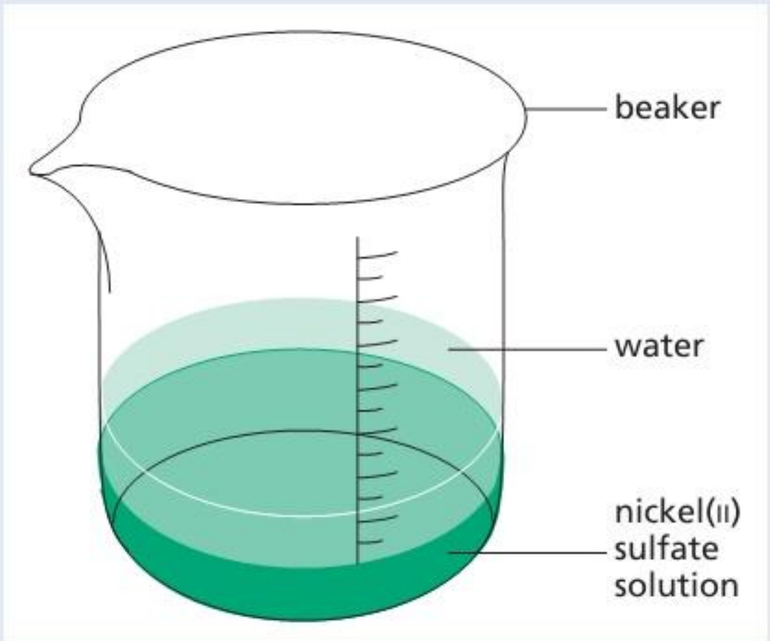
Give explanations for the following observations.
a The formation of a white cloud.
b It took a few minutes before the white cloud formed.
c The white cloud formed further from the cotton wool soaked in ammonia.
d Cooling the concentrated ammonia and hydrochloric acid before carrying out the experiment 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 20×10^6 of them would stretch over a length of only 1 cm.

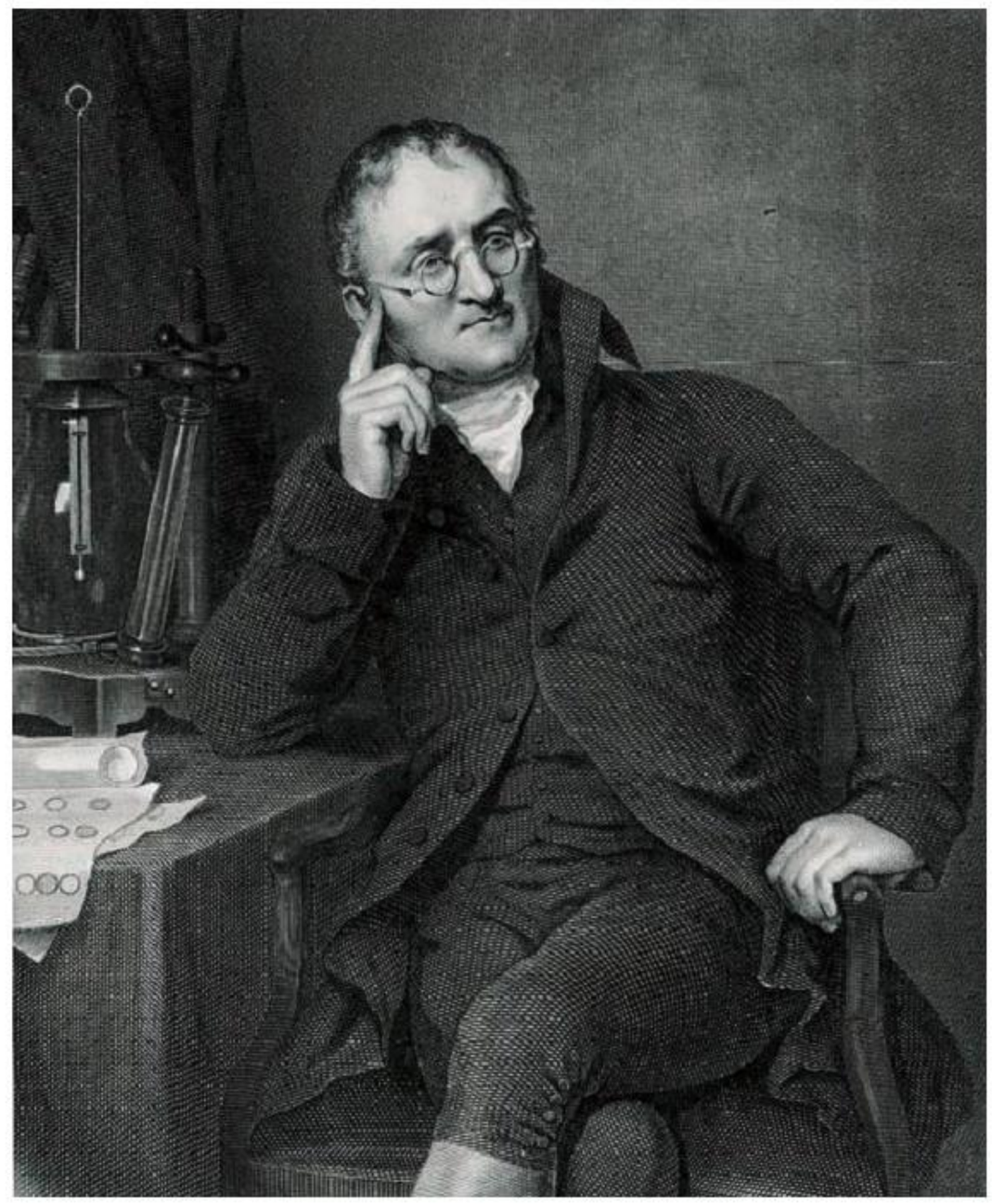


Figure 2.2 John Dalton (1766–1844).

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/ $g\,cm^{-3}$	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[a]	–259	–253
Nitrogen	Non-metal	0.88[b]	–210	–196
Oxygen	Non-metal	1.15[c]	–218	–183
Sulfur	Non-metal	2.07	113	445

Source: Earl B., Wilford L.D.R. Chemistry data book. Nelson Blackie, 1991 [a] At –254°C [b] At –197°C [c] 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 7×10^{-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	7×10^{-8}
Oxygen	12×10^{-8}
Sulfur	20.8×10^{-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.

Cl—Cl

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

1 How would you use a similar chemical shorthand to write a representation of the molecules of iodine and fluorine?

2 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:

hydrogen + oxygen → water

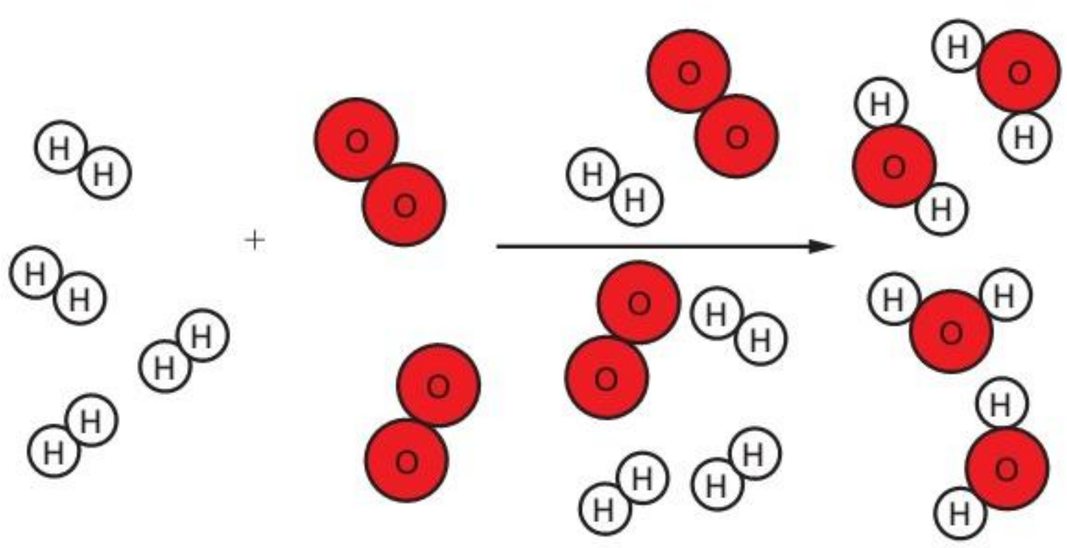


Figure 2.6 The element hydrogen reacts with the element oxygen to produce the compound water.

Water molecules contain two atoms of hydrogen and one atom of oxygen, and hence water has the **chemical formula** H_2O. 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.

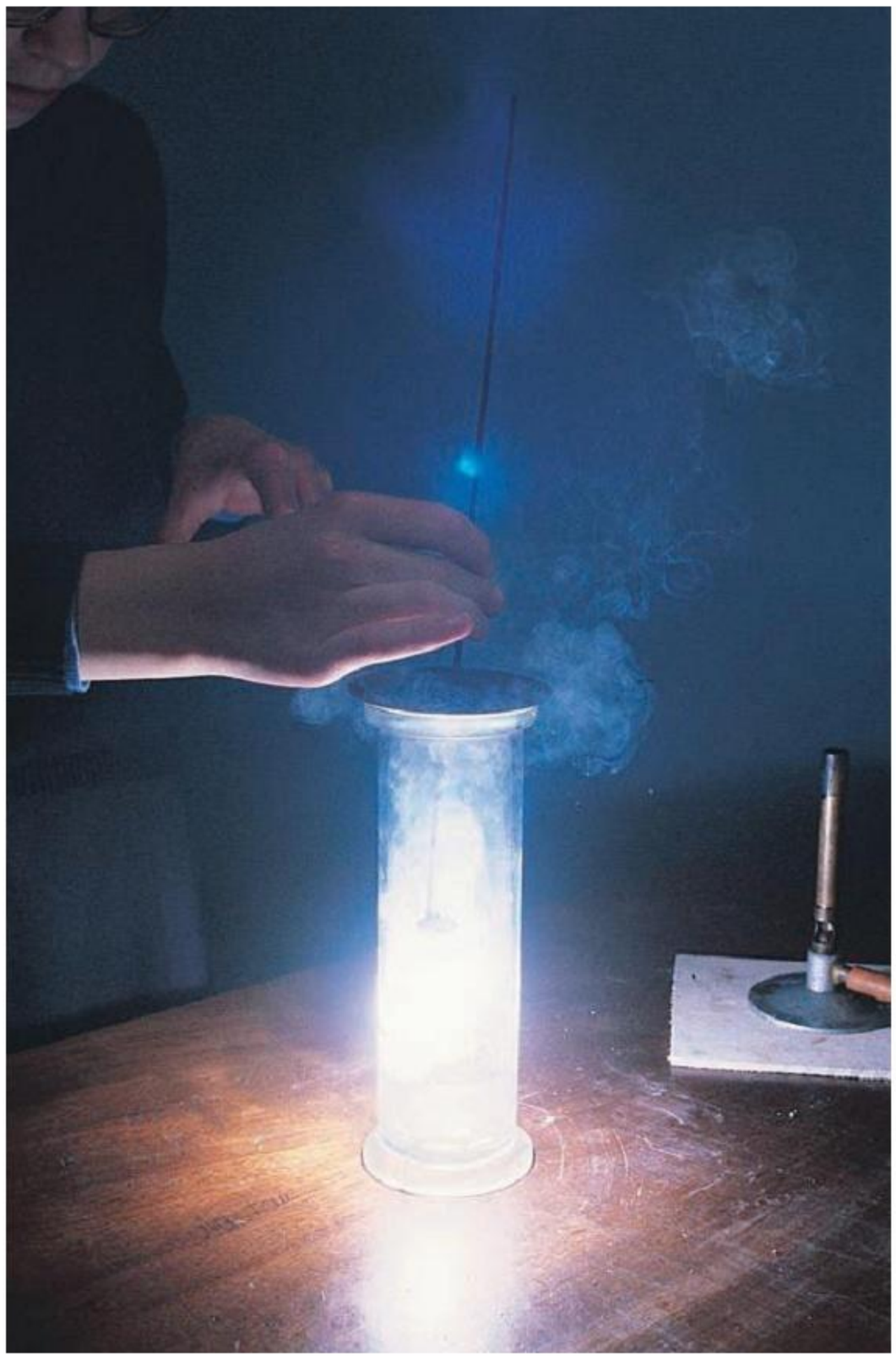


Figure 2.7 Magnesium burns brightly in oxygen to produce magnesium oxide.

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

magnesium + oxygen → magnesium oxide

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.

iron(III) oxide + carbon monoxide → iron + carbon dioxide

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 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	NH_3
Calcium hydroxide	$Ca(OH)_2$
Carbon dioxide	CO_2
Copper sulfate	$CuSO_4$
Ethanol (alcohol)	C_2H_5OH
Glucose	$C_6H_{12}O_6$
Hydrochloric acid	HCl
Nitric acid	HNO_3
Sodium carbonate	Na_2CO_3
Sodium hydroxide	$NaOH$
Sulfuric acid	H_2SO_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:

$$\text{iron} + \text{sulfur} \xrightarrow{\text{heat}} \text{iron(II) sulfide}$$

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

$$Fe(s) + S(s) \xrightarrow{\text{heat}} FeS(s)$$

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:

$$\text{magnesium} + \text{oxygen} \xrightarrow{\text{heat}} \text{magnesium oxide}$$

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:

$$Mg(s) + O_2(g) \xrightarrow{\text{heat}} MgO(s)$$

In the equation there are two oxygen atoms on the left-hand side (O_2) but only one on the right (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 2MgO – this means 2 × MgO. The equation now becomes:

$$Mg(s) + O_2(g) \xrightarrow{\text{heat}} 2MgO(s)$$

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:

$$2Mg(s) + O_2(g) \xrightarrow{\text{heat}} 2MgO(s)$$

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_2, N_2 or H_2 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

1 Write the word and balanced chemical equations for the reactions which take place between:
   a calcium and oxygen b copper and oxygen.
2 Write down the ratio of the atoms present in the formula for each of the compounds shown in Table 2.5.
3 Iron is extracted from iron(III) oxide in a blast furnace by a redox reaction. What does the term 'redox reaction' mean?
4 Identify the oxidising and reducing agents in the following reactions:
   a copper(II) oxide + hydrogen → copper + water
   b tin(II) oxide + carbon → tin + carbon dioxide
   c $PbO(s) + H_2(g) \rightarrow Pb(s) + H_2O(l)$

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:

$$\text{iron} + \text{sulfur} \xrightarrow{\text{heat}} \text{iron(II) sulfide}$$

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.

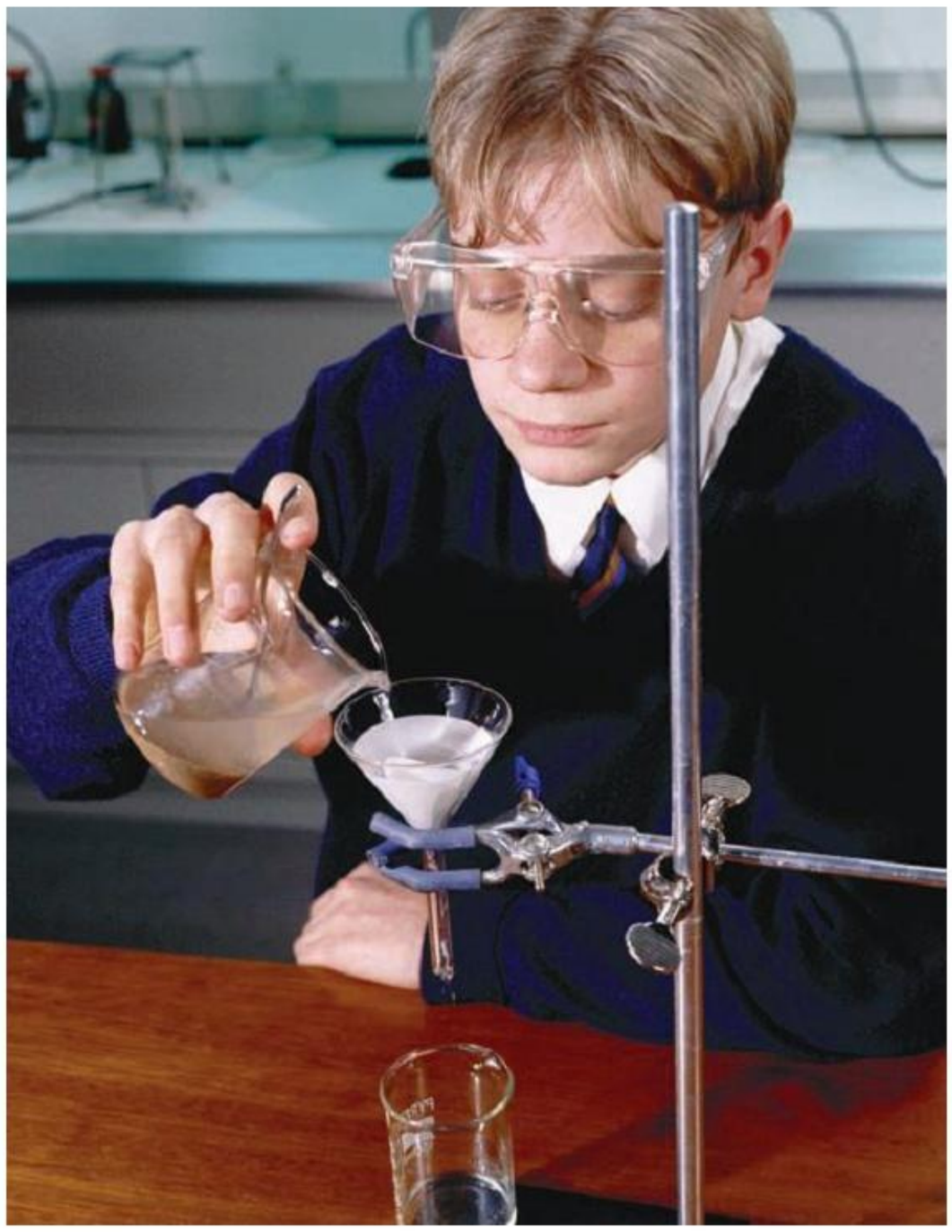


Figure 2.13 It is important when filtering not to overfill the filter paper.

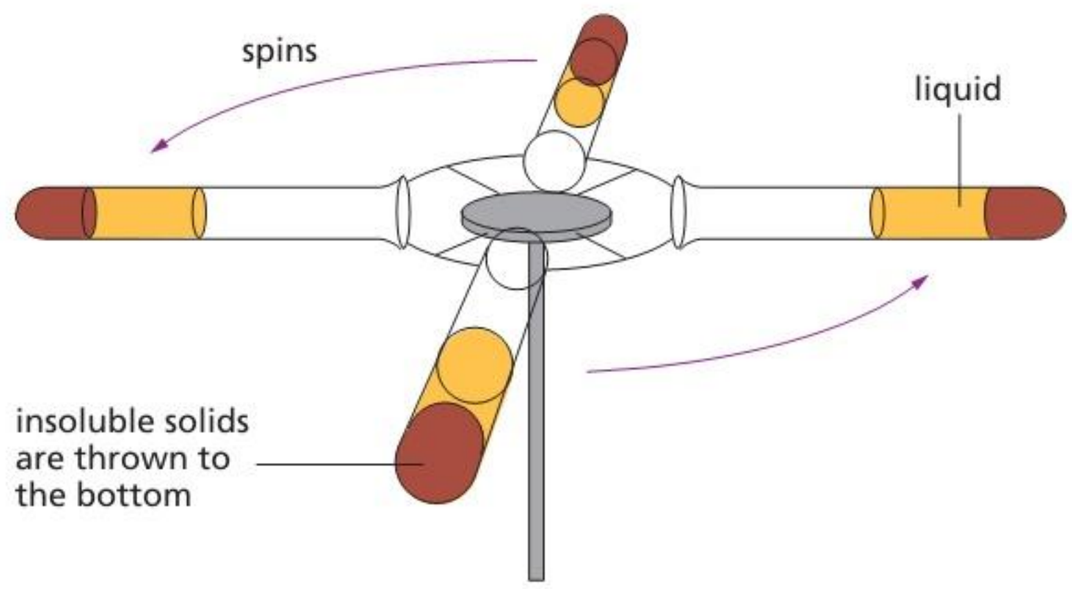
The filter paper contains holes that, although too small to be seen, are large enough to allow the molecules of water through but not the sand particles. It acts like a sieve. The sand gets trapped in the filter paper and the water passes through it. The sand is called the **residue** and the water is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

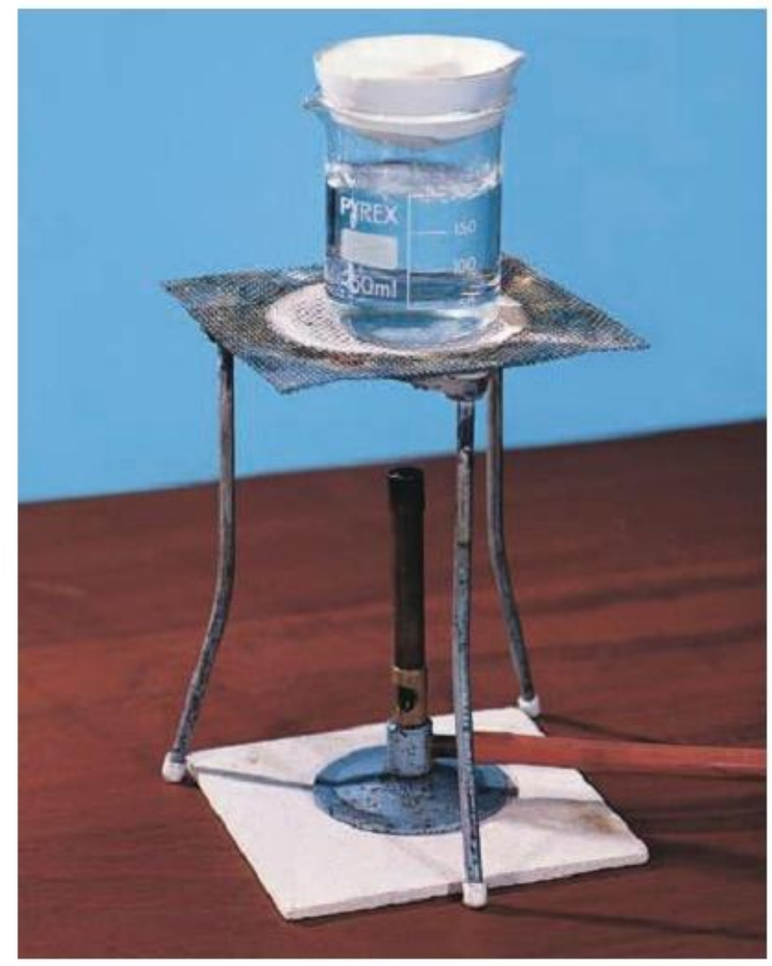


Figure 2.16 Apparatus used to slowly evaporate a solvent.

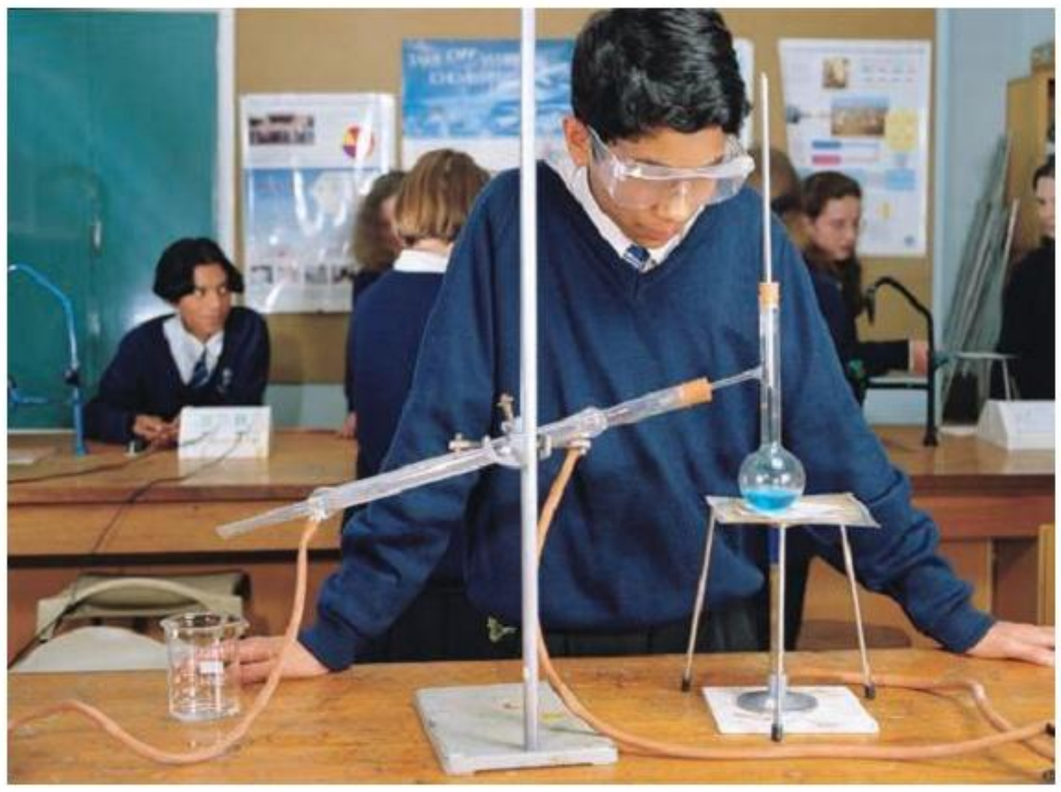
Crystallisation

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.



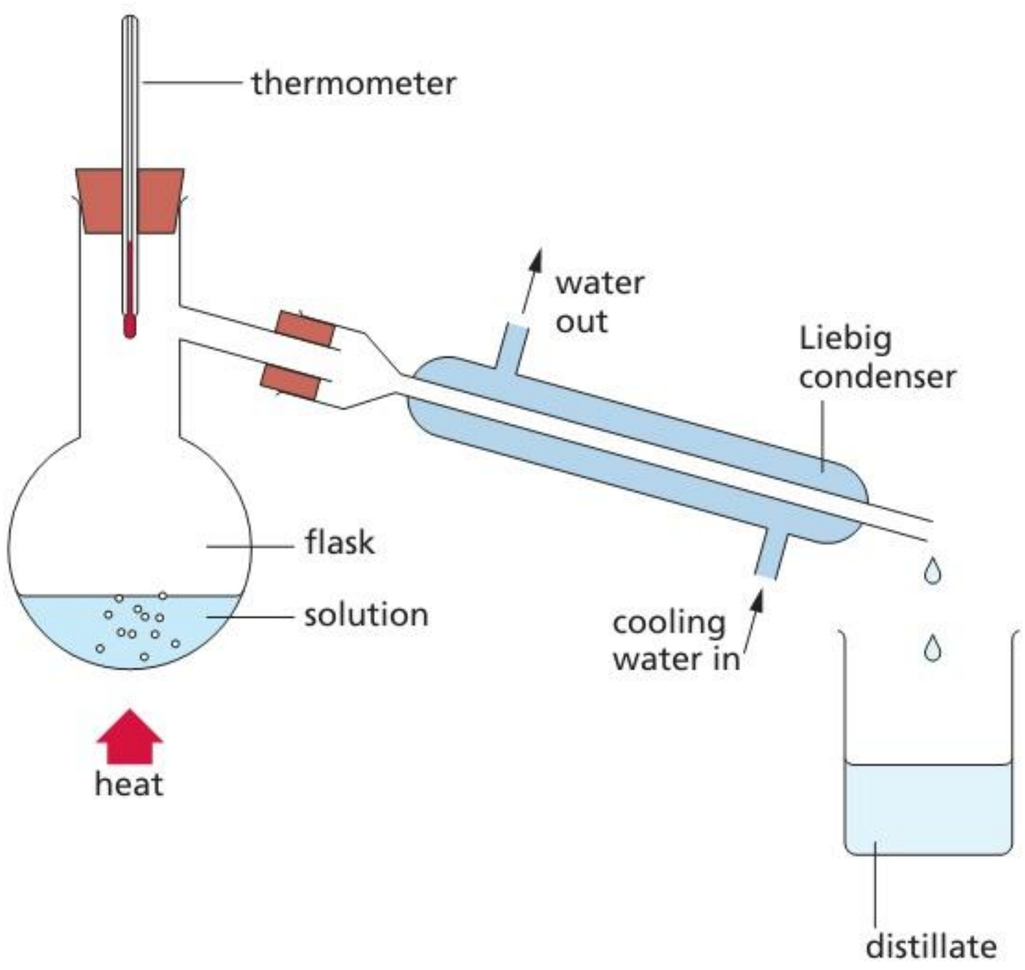


Figure 2.18 Water can be obtained from salt water by distillation.

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 °C whereas water boils at 100 °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 steadily to 100 °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.

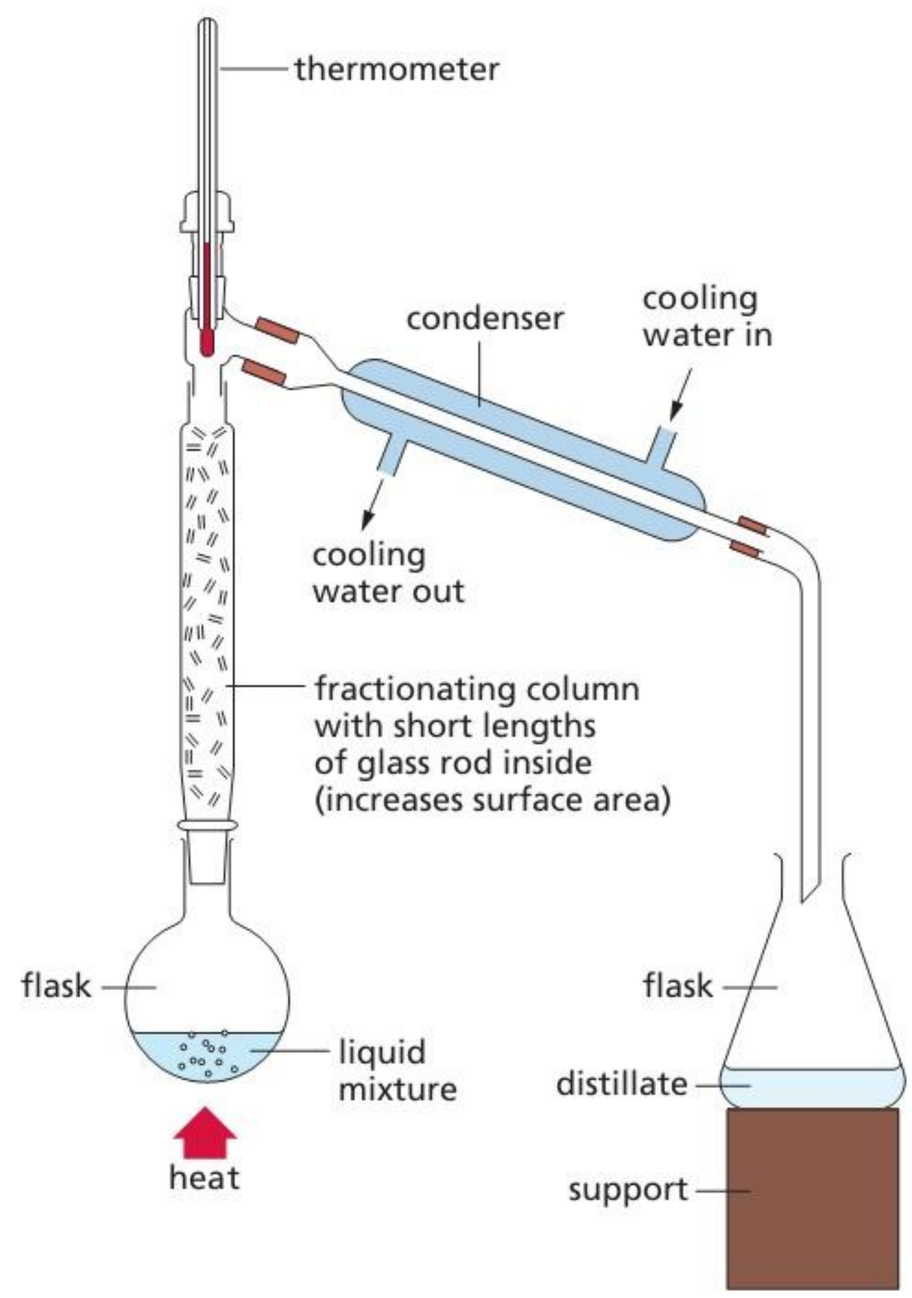


Figure 2.22 Typical fractional distillation apparatus.

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.

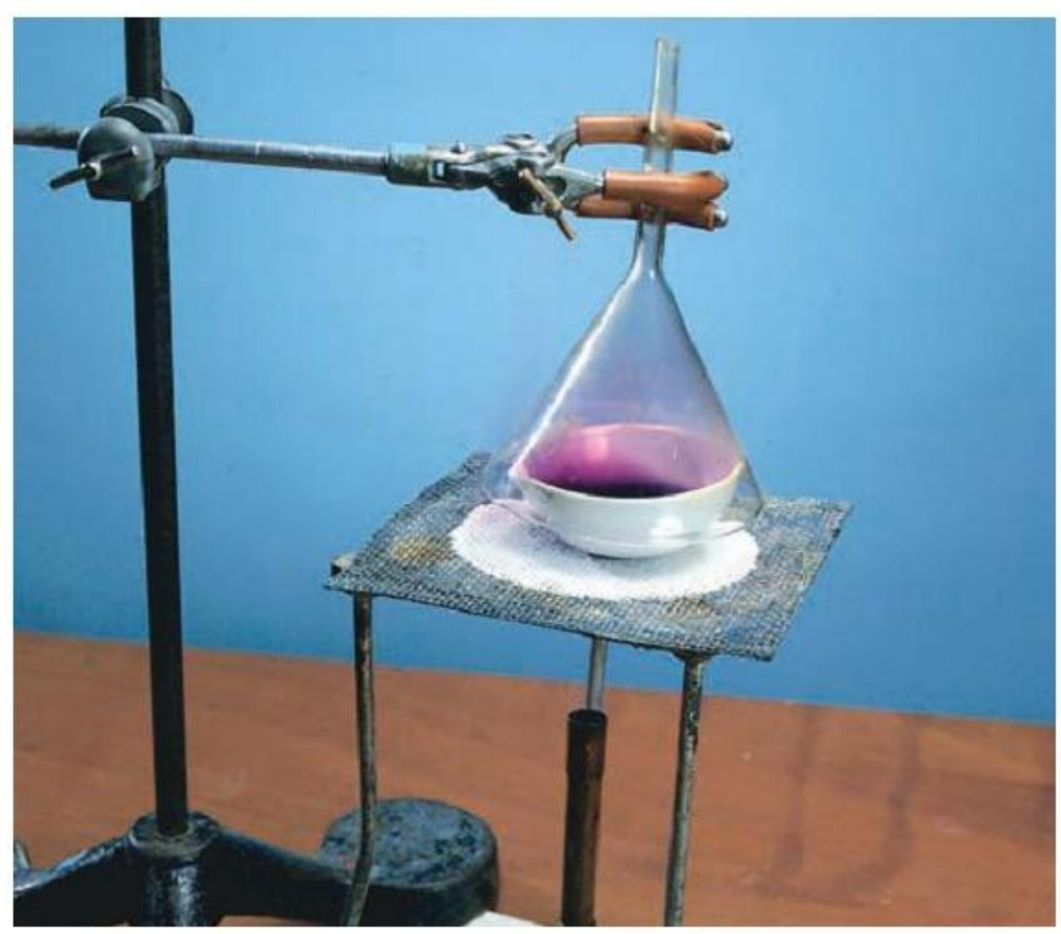


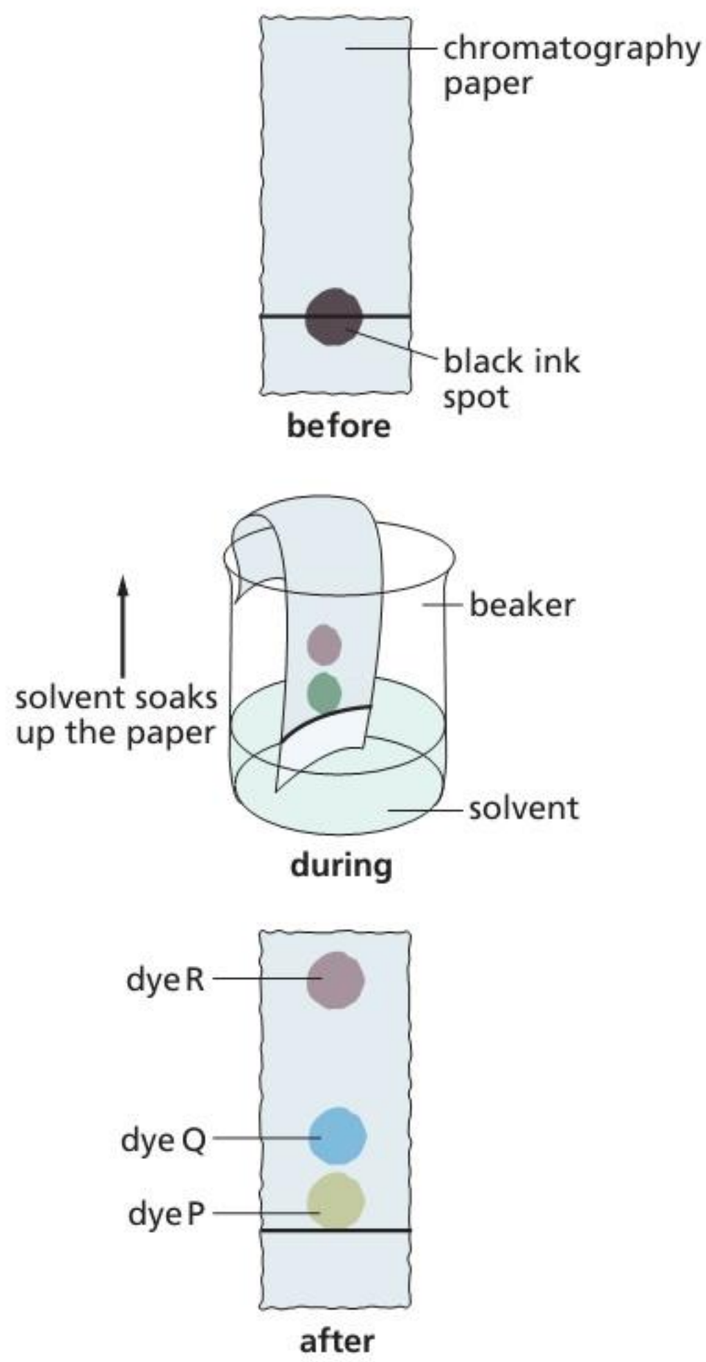
Figure 2.25 Apparatus used to separate an iodine/salt mixture. The iodine sublimes on heating.

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).

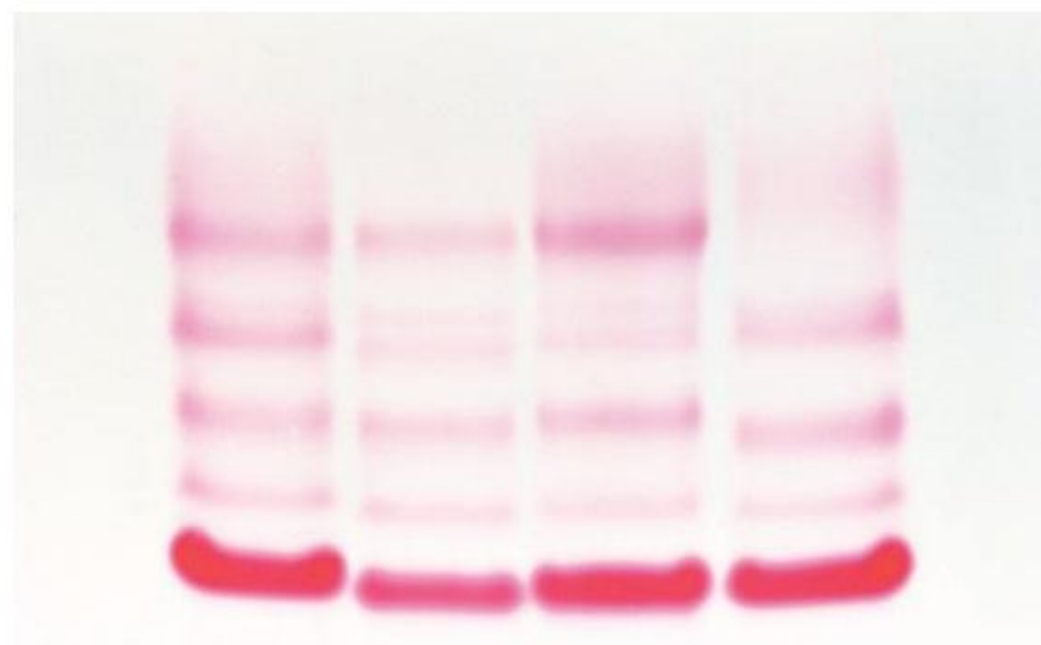


Figure 2.27 Protein samples are separated by electrophoresis in medical research laboratories.

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.

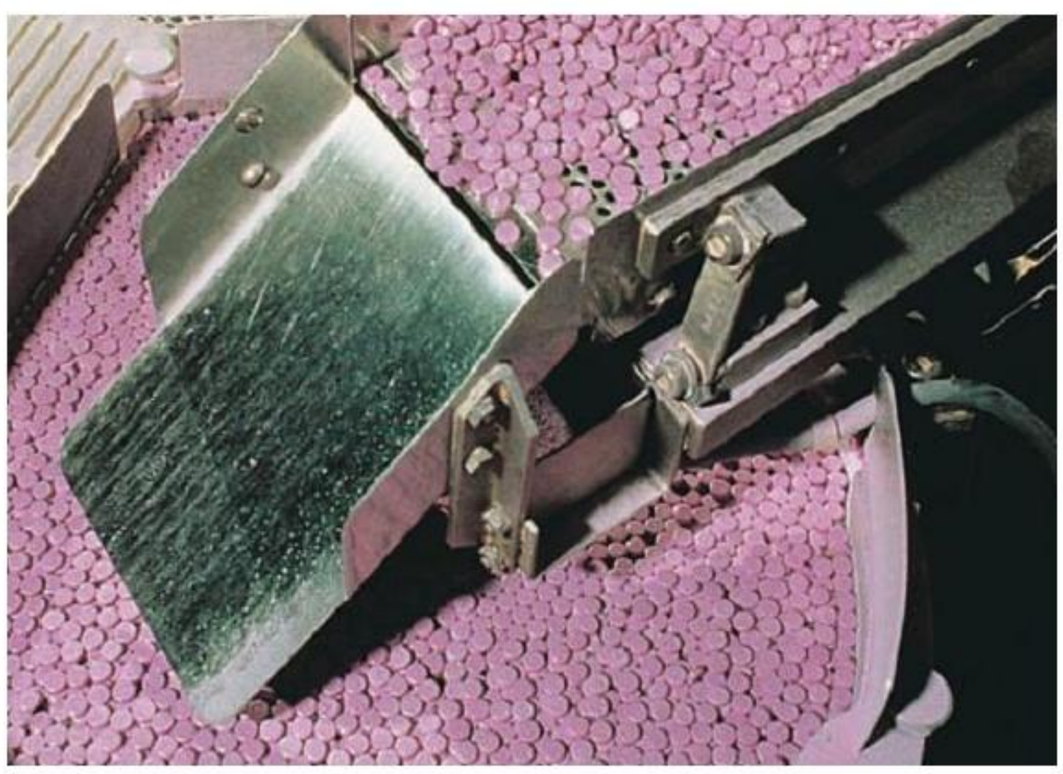


Figure 2.29 Drugs are manufactured to a high degree of purity by fractional crystallisation.

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

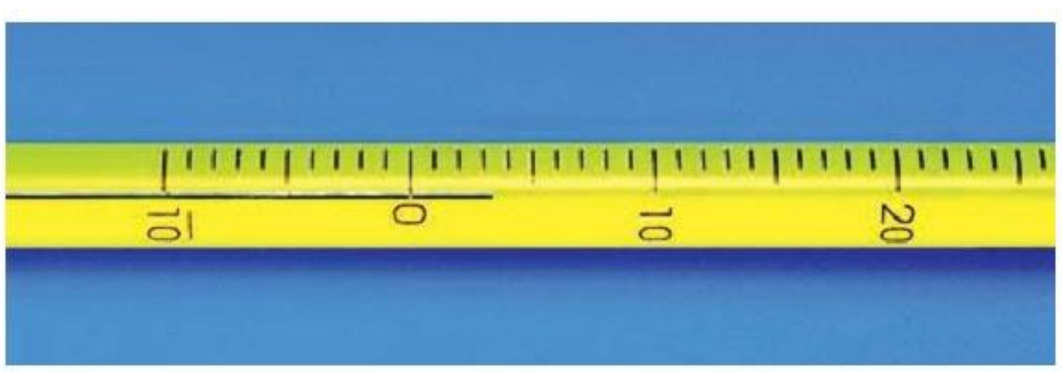


Figure 2.32 A thermometer can be used to measure temperature.

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 °C whilst
the boiling point of water is 100 °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° and 110°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 kg = 1000 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^3).

1 litre = 1000 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^3 = 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^{-9} 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).

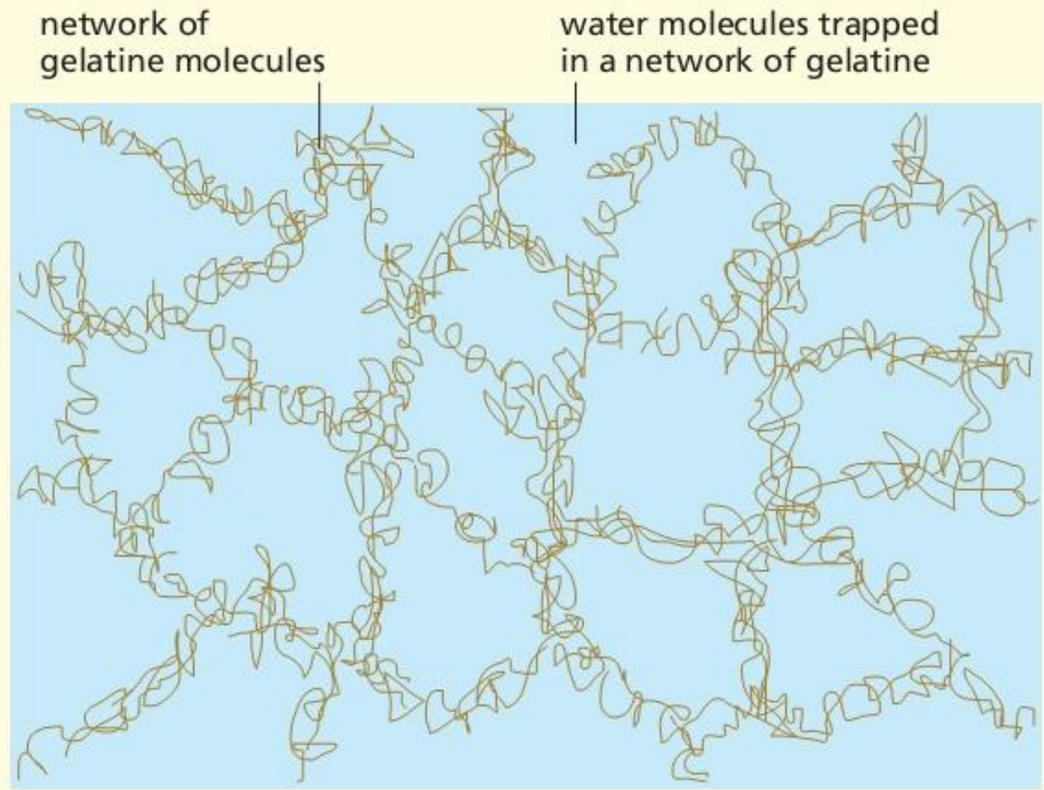


Figure 2.36 The network within a gel.

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.

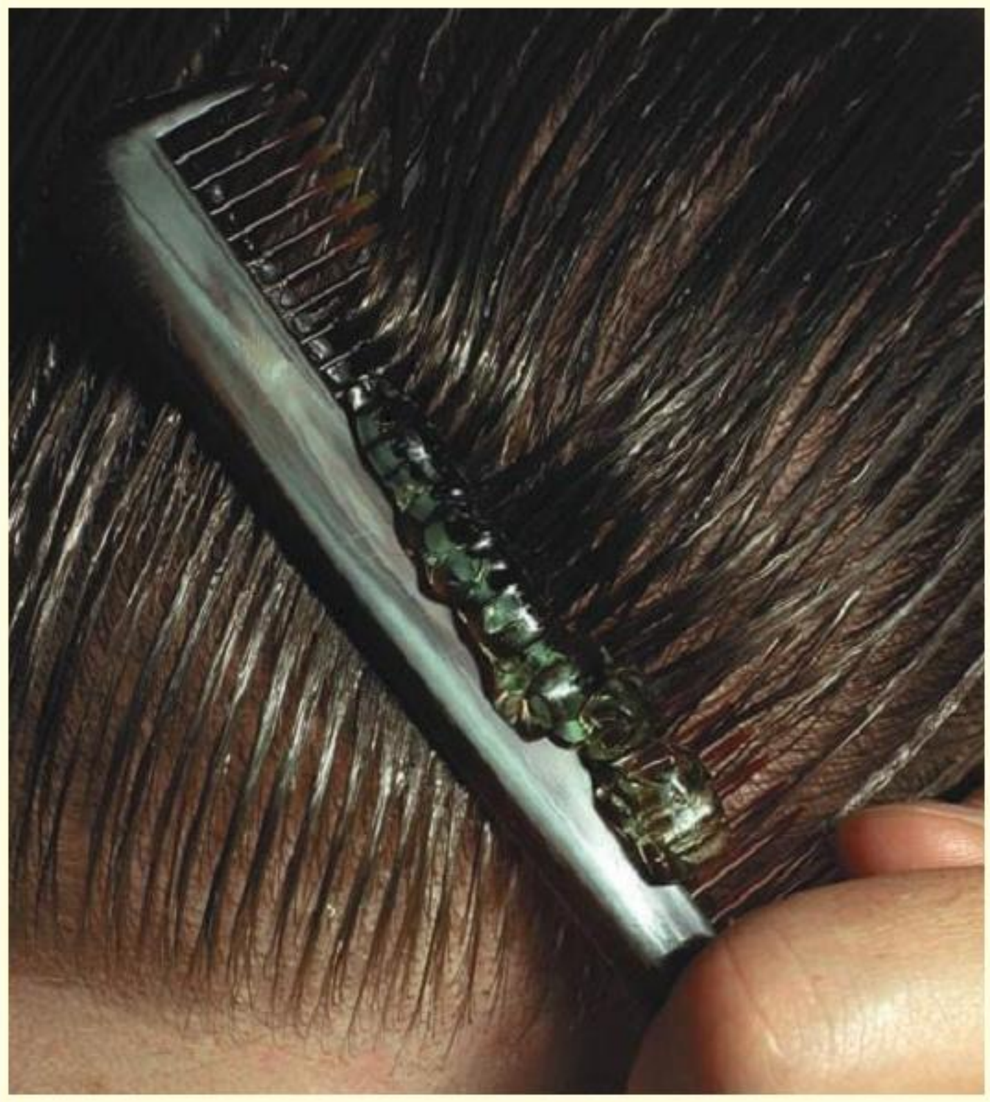


Figure 2.37 Hair gel is a mixture of water and an oil plus a perfume.

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

1 Explain the following terms:
   a colloid
   b emulsifier
   c foam
   d 'E' number
   e sol.

2 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).

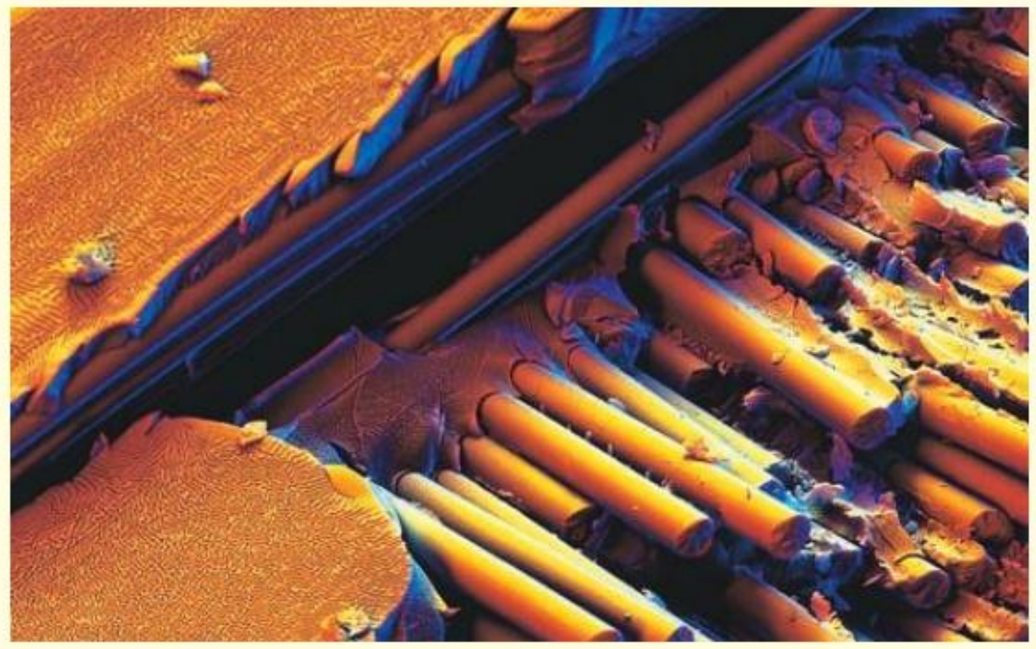


Figure 2.40 GRP consists of glass fibres (rod shapes) embedded in plastic, in this case polyester.

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.

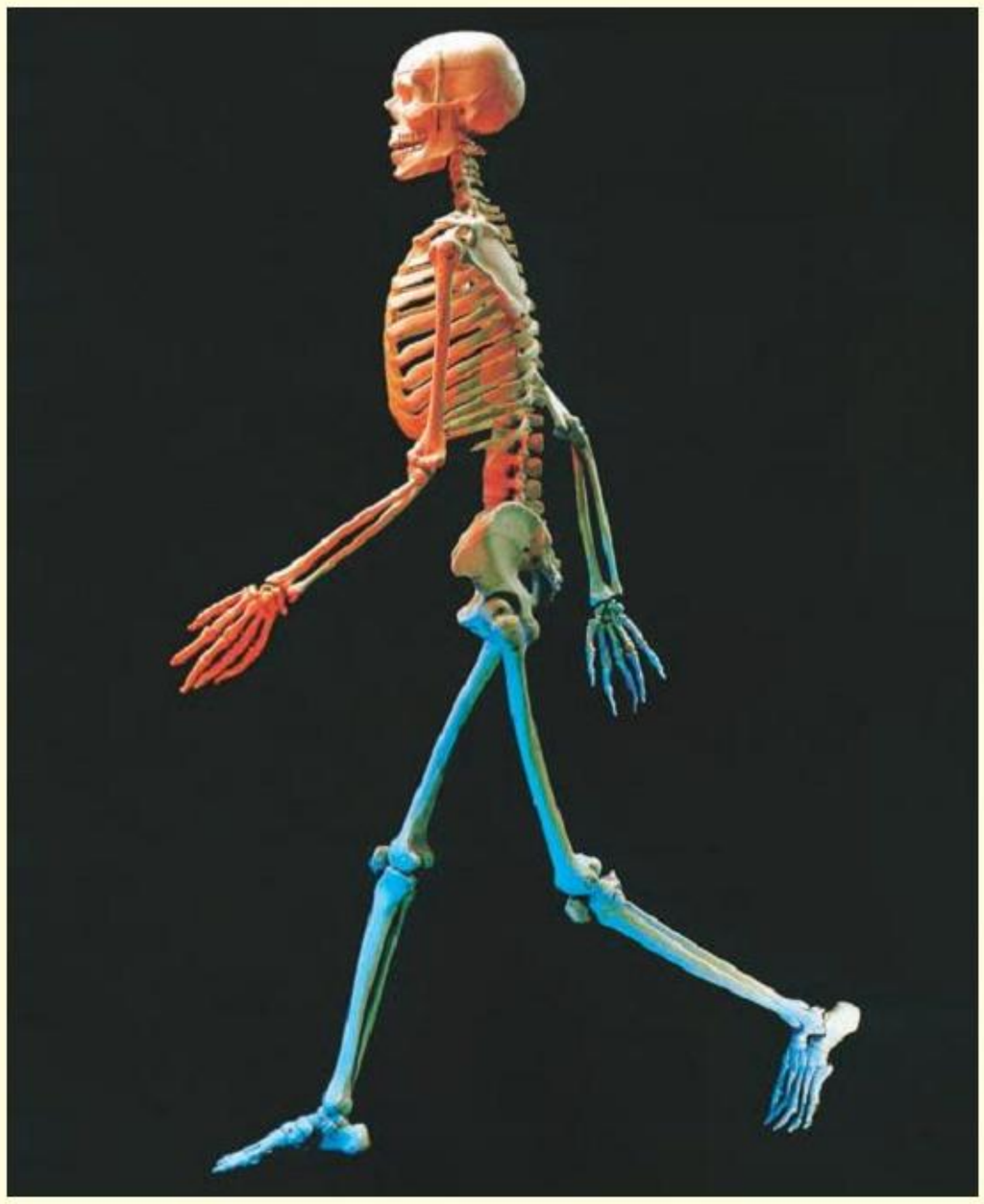


Figure 2.42 Bone is a composite material.

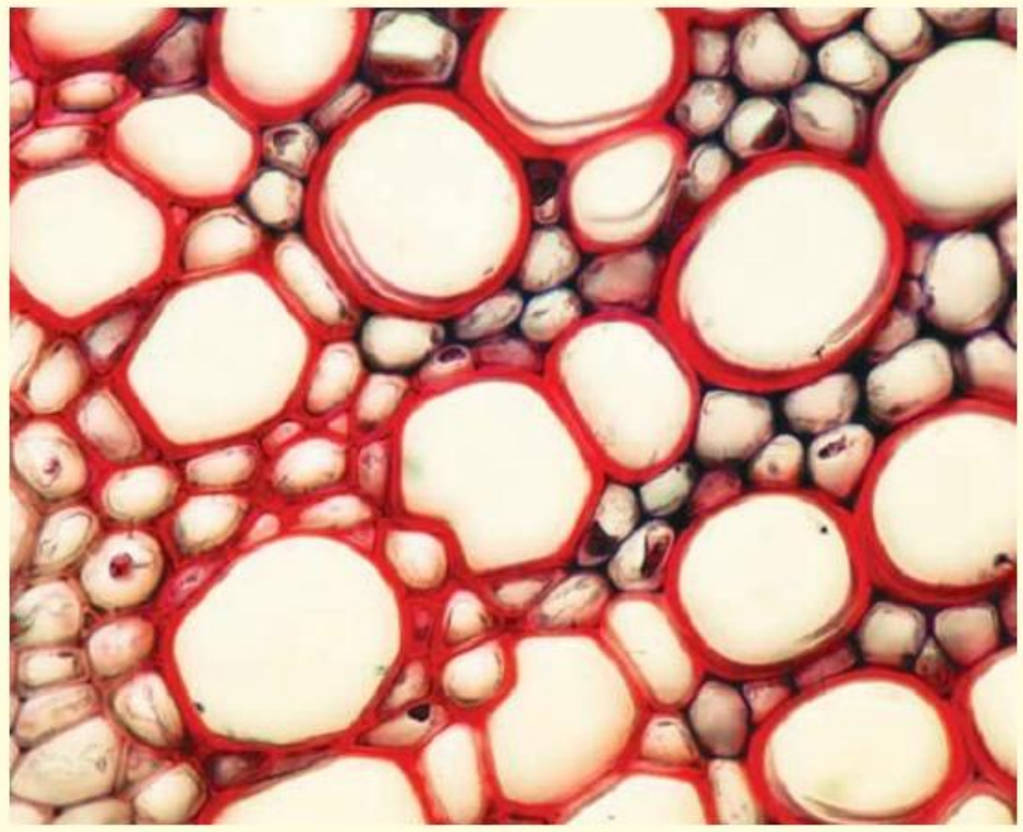


Figure 2.43 The combination of cellulose fibres and lignin makes the cell walls hard, thick and very strong. These properties reinforce the cells against collapse.

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 nm = 1×10^{-9} 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, H_2, and oxygen, 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_f 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.

$$\text{solute} + \text{solvent} \xrightarrow{\text{dissolves}} \text{solution}$$

Elements, compounds and experimental techniques

Additional questions

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

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

2 Which of the substances listed below are:

a metallic elements?
b non-metallic elements?
c compounds?
d 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 a solid element?
b a liquid element?
c a gaseous mixture?
d a solid mixture?
e a liquid compound?
f 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.

a Explain what the red glow indicates.
b Give the chemical name of the black solid.
c Write a word equation and a balanced chemical equation to represent the reaction which has taken place.
d 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:

a oxygen from liquid air
b red blood cells from plasma
c petrol and kerosene from crude oil
d coffee grains from coffee solution
e pieces of steel from engine oil
f amino acids from fruit juice solution
g 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^{-3}
A	1110	2606	9.1
B	−266	−252	0.07
C	40	94	1.6
D	−14	60	0.9

a Which substance is a gas at room temperature?
b Which substance is a liquid at room temperature?
c Which substances are solids at room temperature?
d Which substance is most likely to be a metal?
e Which substance will be a liquid at −260 °C?
f What is the melting point of the least dense non-metal?
g 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?

(i) nitric acid, HNO_3
(ii) methane, CH_4
(iii) copper nitrate, $Cu(NO_3)_2$
(iv) ethanoic acid, CH_3COOH
(v) sugar, $C_{12}H_{22}O_{11}$
(vi) phenol, C_6H_5OH
(vii) ammonium sulfate, $(NH_4)_2SO_4$

b Balance the following equations:

(i) $Zn(s) + O_2(g) \rightarrow ZnO(s)$
(ii) $Fe(s) + Cl_2(g) \rightarrow FeCl_3(s)$
(iii) $Li(s) + O_2(g) \rightarrow Li_2O(s)$
(iv) $H_2(g) + O_2(g) \rightarrow H_2O(g)$
(v) $Mg(s) + CO_2(g) \rightarrow MgO(s) + C(s)$

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^{-3}	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.

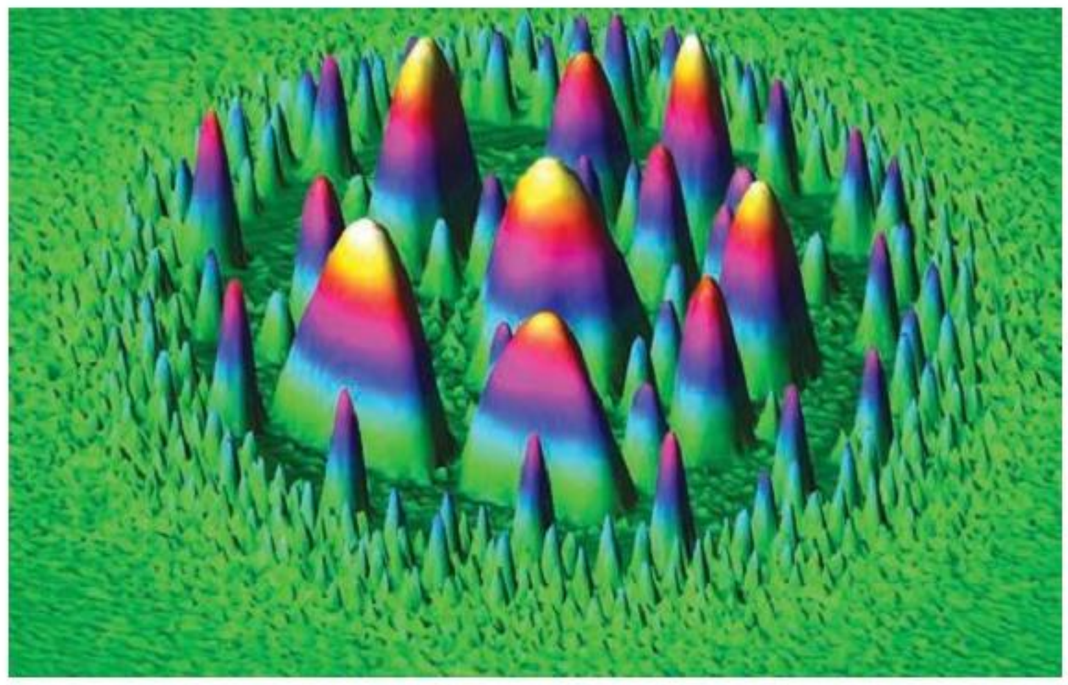


Figure 3.1 A micrograph of atoms.

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).

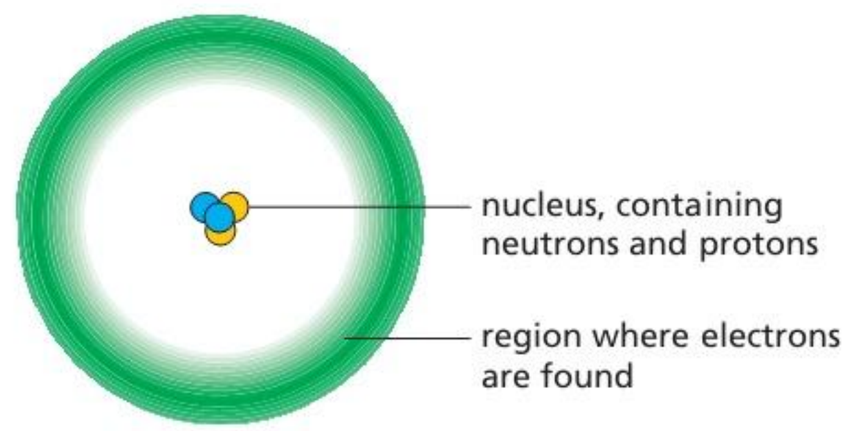


Figure 3.2 Diagram of an atom.

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.

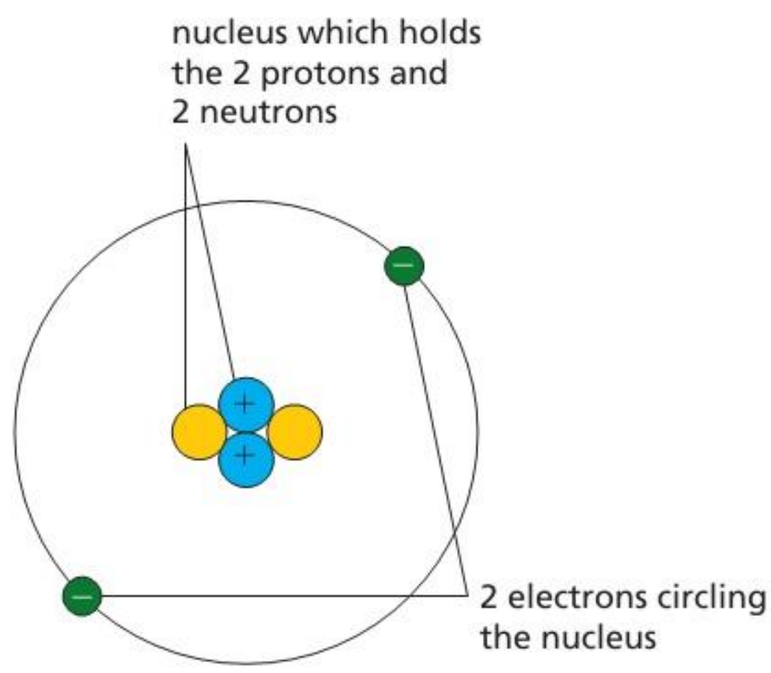


Figure 3.3 An atom of helium has two protons, two electrons and two neutrons.

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***.

$$\underset{(A)}{\text{nucleon number}} = \underset{(Z)}{\text{proton number}} + \underset{\text{neutrons}}{\text{number of}}$$

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:

nucleon number (A) → $^{4}_{2}\text{He}$ ← symbol of the element
proton number (Z) ↗

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

$$\underset{\text{neutrons}}{\text{number of}} = \underset{(A)}{\text{nucleon number}} - \underset{(Z)}{\text{proton number}}$$

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

$$\underset{(A)}{24} - \underset{(Z)}{12} = 12$$

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

$$\underset{(A)}{207} - \underset{(Z)}{82} = 125$$

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, K^+.

	19 protons	= 19+
$_{19}K^+$	18 electrons	= 18−
	Overall charge = 1+	

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, O^{2-}.

	8 protons	= 8+
$_{8}O^{2-}$	10 electrons	= 10−
	Overall charge = 2−	

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 NO_3^-
- an ion may possess more than one unit of charge (either negative or positive), for example Al^{3+}, O^{2-} or SO_4^{2-}.

Table 3.3 Some common ions.

Name	Formula
Lithium ion	Li^+
Sodium ion	Na^+
Potassium ion	K^+
Magnesium ion	Mg^{2+}
Calcium ion	Ca^{2+}
Aluminium ion	Al^{3+}
Zinc ion	Zn^{2+}
Ammonium ion	NH_4^+
Fluoride ion	F^-
Chloride ion	Cl^-
Bromide ion	Br^-
Hydroxide ion	OH^-
Oxide ion	O^{2-}
Sulfide ion	S^{2-}
Carbonate ion	CO_3^{2-}
Nitrate ion	NO_3^-
Sulfate ion	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.

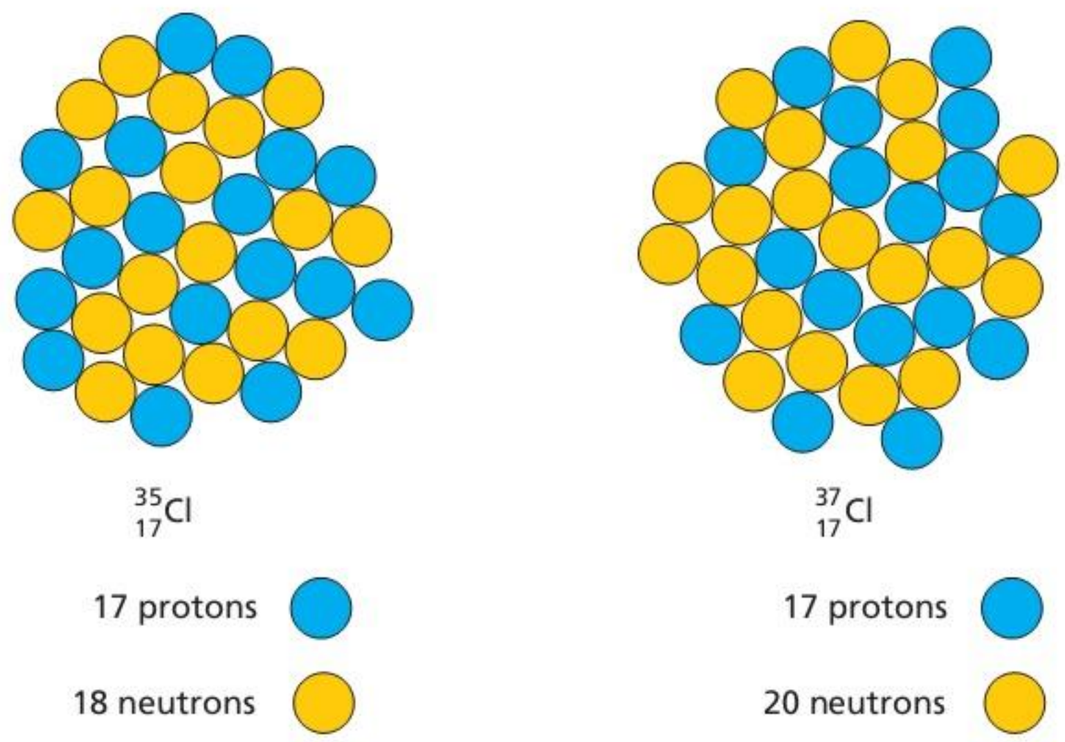


Figure 3.4 The two isotopes of chlorine.

Table 3.4 Some atoms and their isotopes.

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

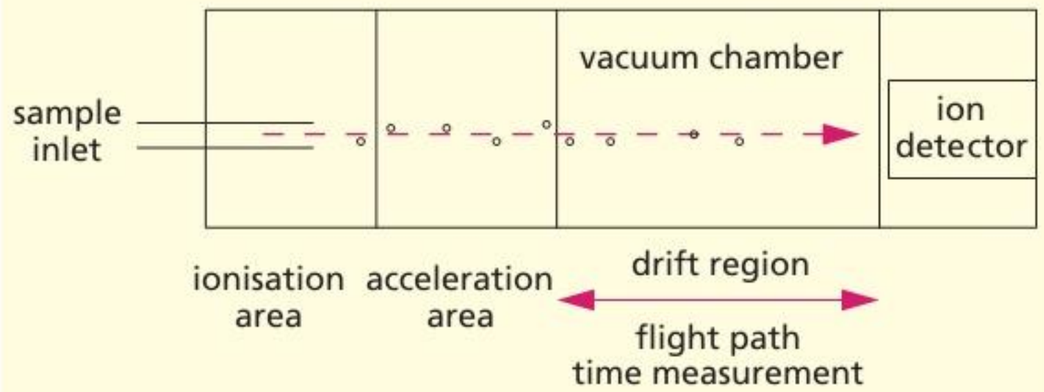


Figure 3.6 A diagram of a mass spectrometer.

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.

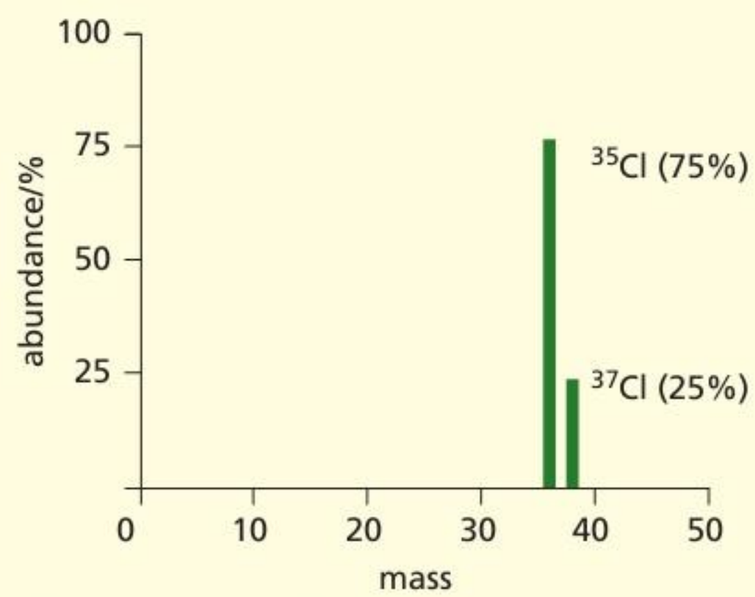


Figure 3.7 The mass spectrum for chlorine.

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}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:

	$^{35}_{17}Cl$	$^{37}_{17}Cl$
% 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

1 Calculate the number of neutrons in the following atoms:
a $^{27}_{13}Al$ b $^{31}_{15}P$ c $^{262}_{107}Uns$ d $^{190}_{76}Os$

2 Given that the percentage abundance of $^{20}_{10}Ne$ is 90% and that of $^{22}_{10}Ne$ 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 $^{16}_{8}O$ 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.

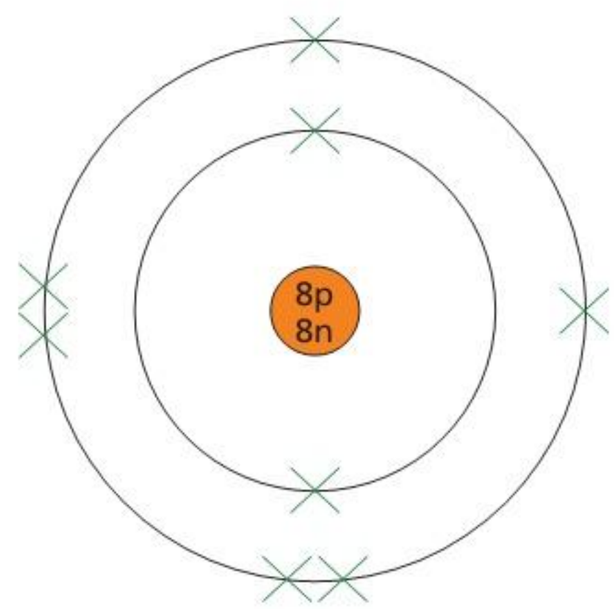


Figure 3.8 Arrangement of electrons in an oxygen atom.

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

1 How many electrons may be accommodated in the first three energy levels?

2 What is the same about the electron structures of:
   a lithium, sodium and potassium?
   b beryllium, magnesium and calcium?

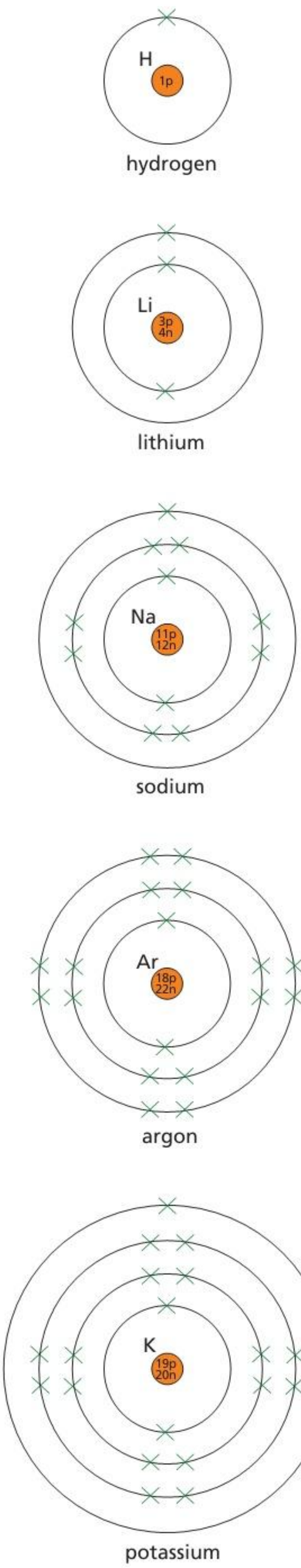


Figure 3.9 Electron arrangements of hydrogen, lithium, sodium, argon and potassium.

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

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 → sodium chloride

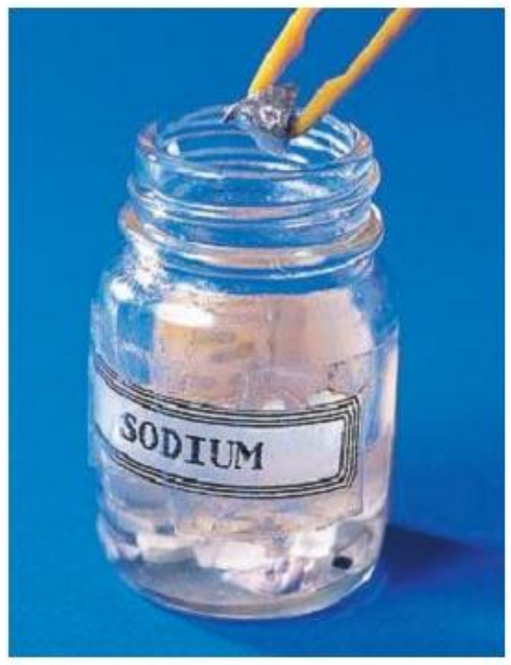


Figure 3.10 The properties of salt are very different from those of the sodium and chlorine it was made from. To get your salt you would not eat sodium or inhale chlorine!

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

sodium atom	→	sodium ion	+	electron
$Na(g)$	→	$Na^+(g)$	+	e^-

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

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

chlorine atom	+	electron	→	chloride ion
$Cl(g)$	+	e^-	→	$Cl^-(g)$

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.

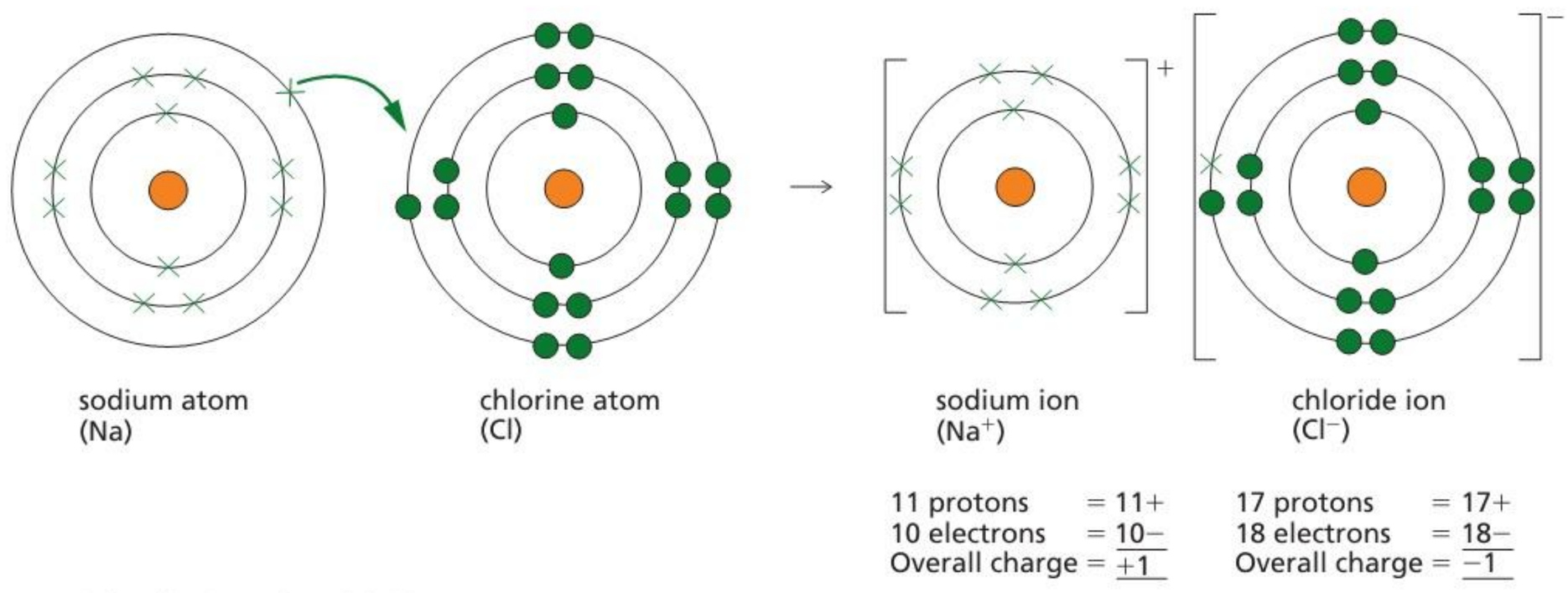


Figure 3.11 Ionic bonding in sodium chloride.

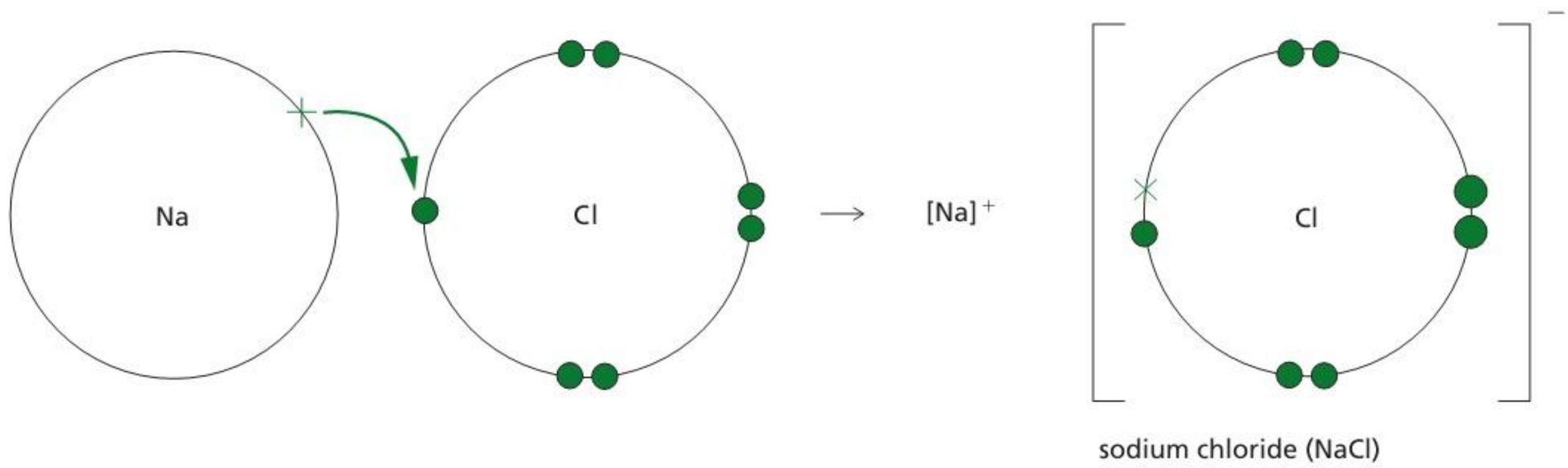


Figure 3.12 Simplified diagram of ionic bonding in sodium chloride.

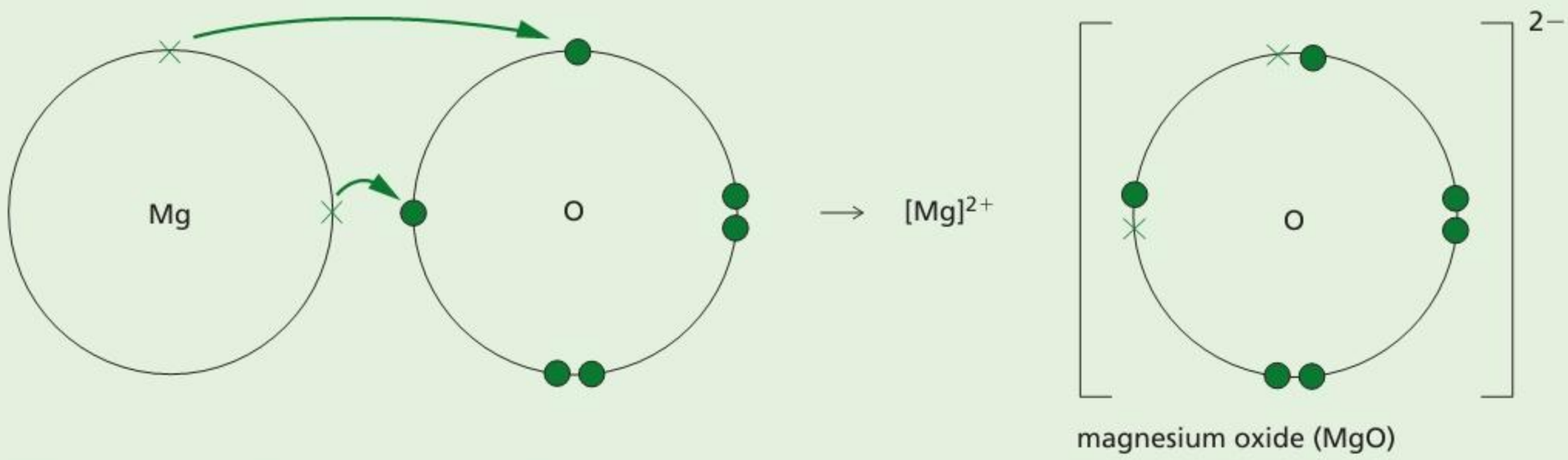


Figure 3.13 Simplified diagram of ionic bonding in magnesium oxide.

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 Mg^{2+} and O^{2-} are oppositely charged and are attracted to one another. The formula for magnesium oxide is 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 (Cl^-) for each calcium ion (Ca^{2+}). The chemical formula is $CaCl_2$.

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 (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.

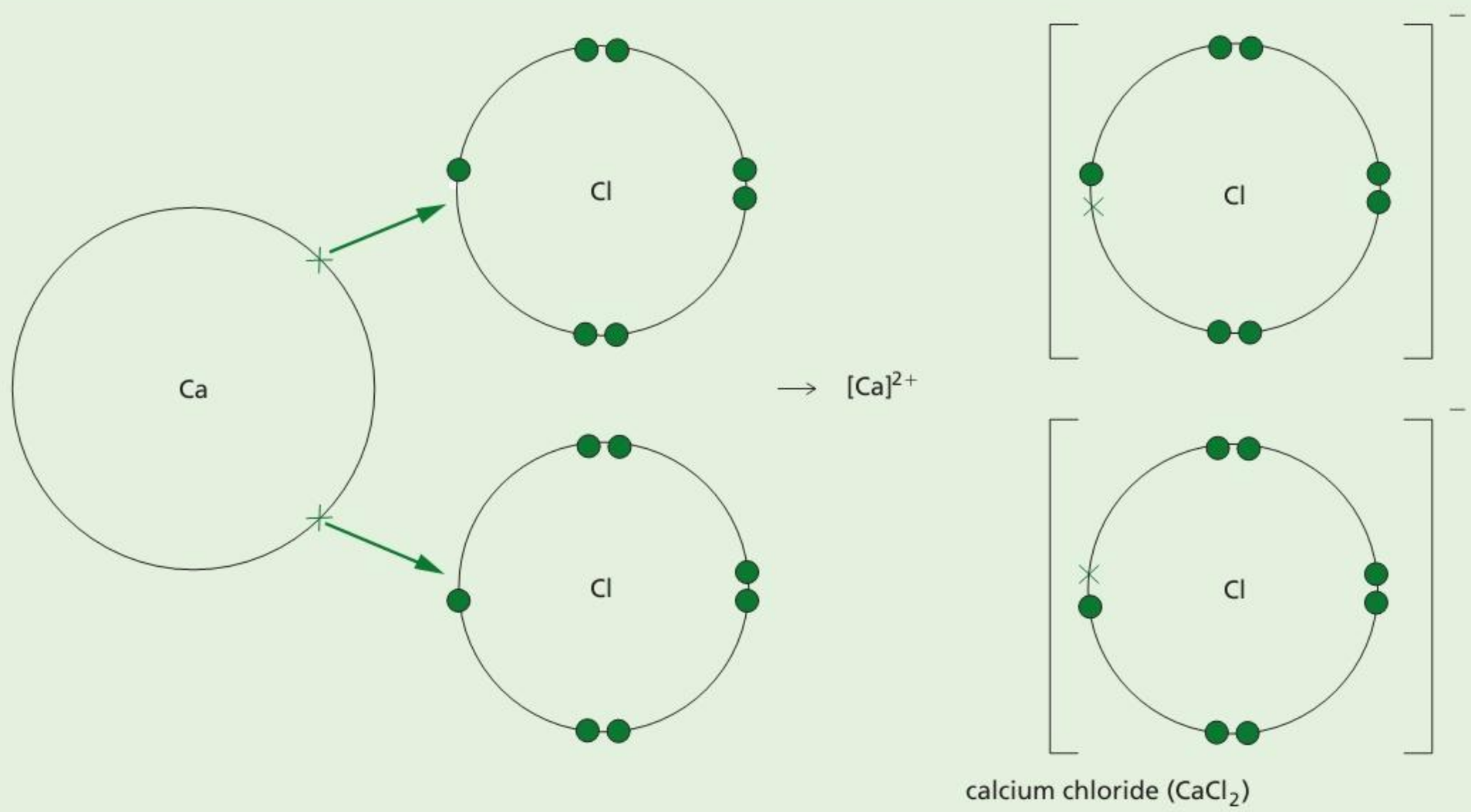
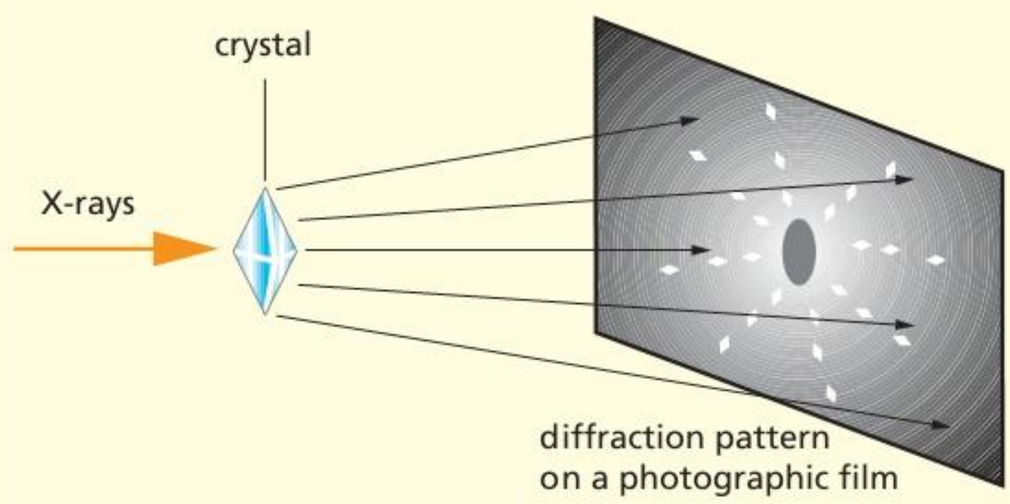
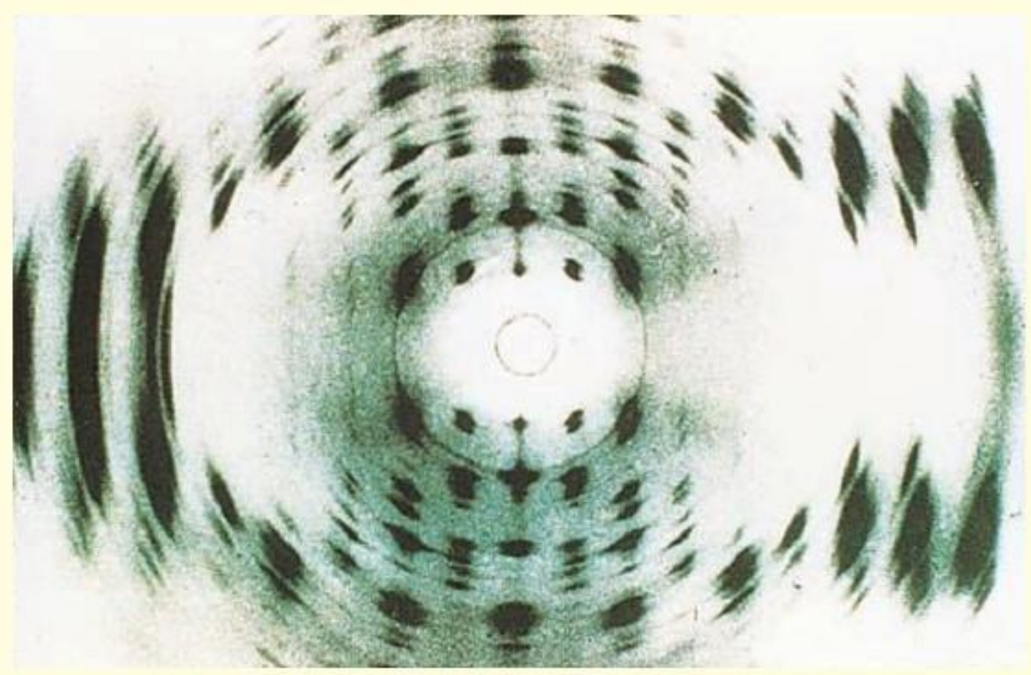


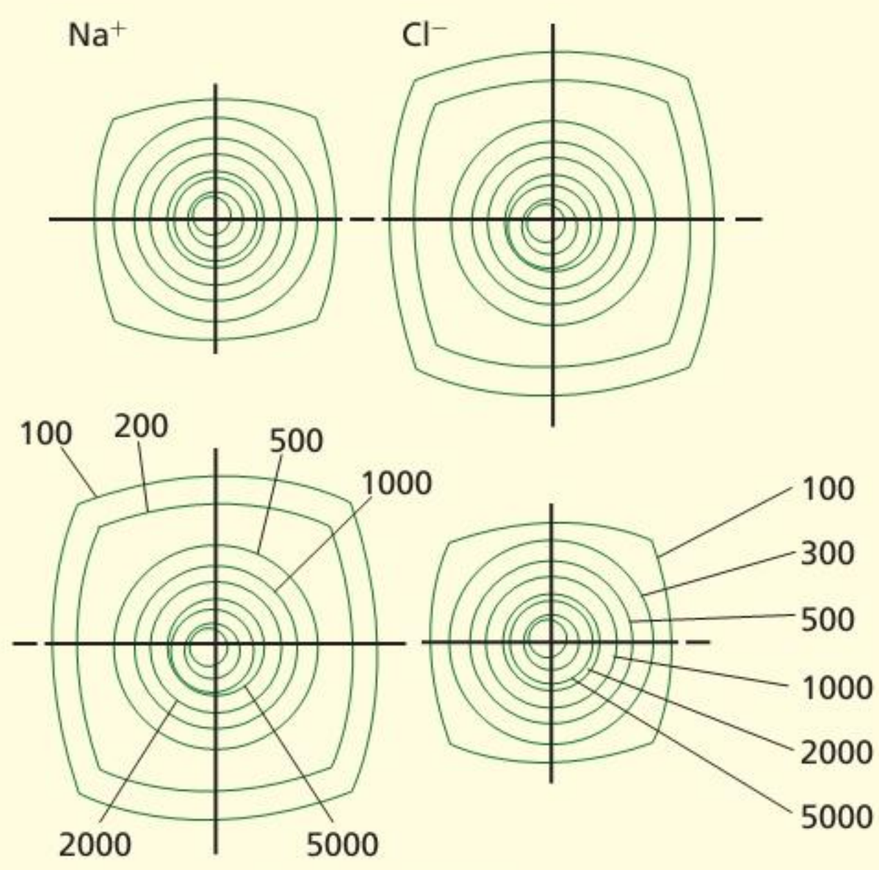
Figure 3.14 The transfer of electrons that takes place during the formation of calcium chloride.



a X-ray diffraction technique.

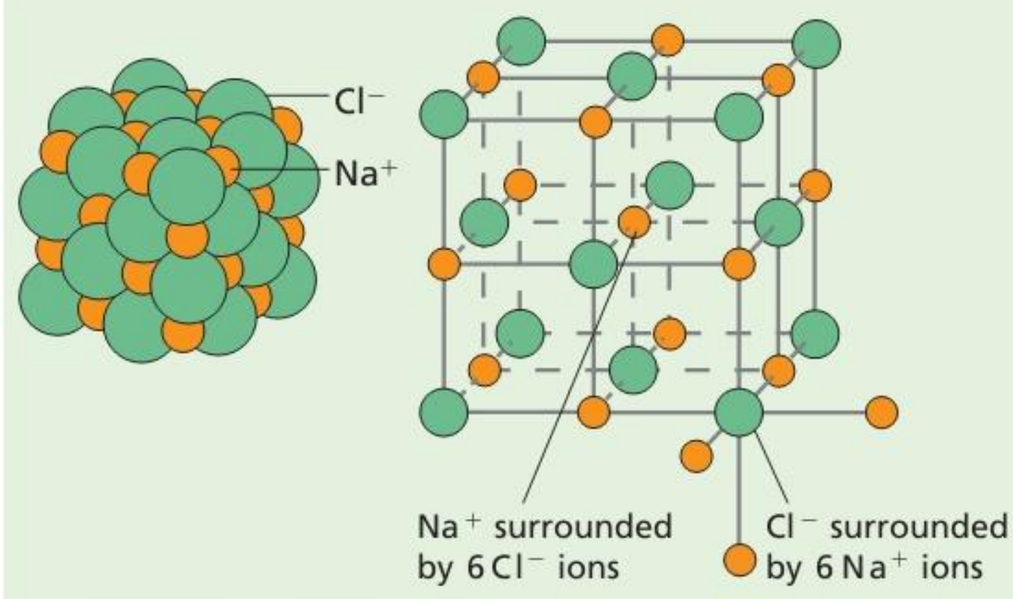


b X-ray diffraction photograph of sodium chloride.



c Electron density map of sodium chloride.

Figure 3.15



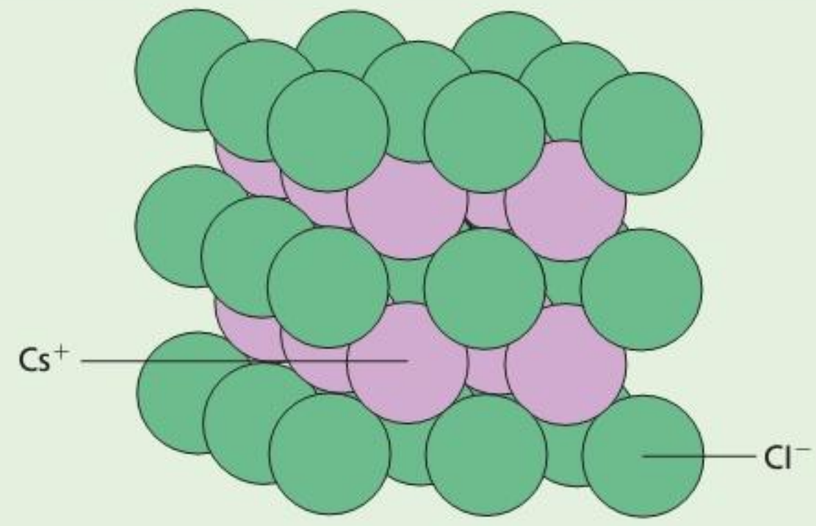
d The structure of sodium chloride.

Ionic structures

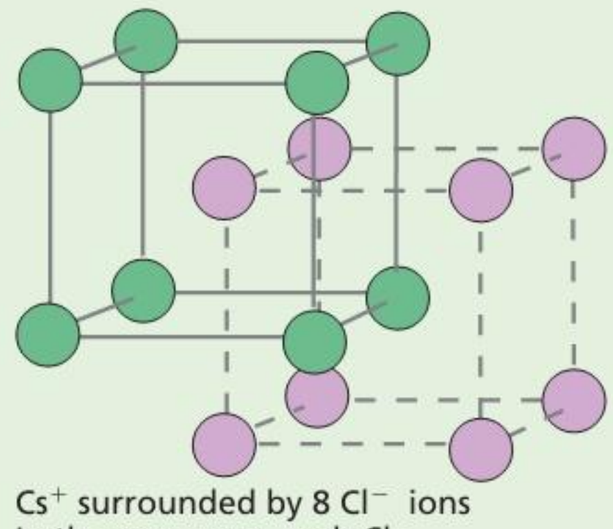
Ionic structures are solids at room temperature and have high melting and boiling points. The ions are packed together in a regular arrangement called a **lattice**. Within the lattice, oppositely charged ions attract one another strongly.

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 (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.



Cs^+ surrounded by 8 Cl^- ions
In the same way, each Cl^- is surrounded by 8 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).

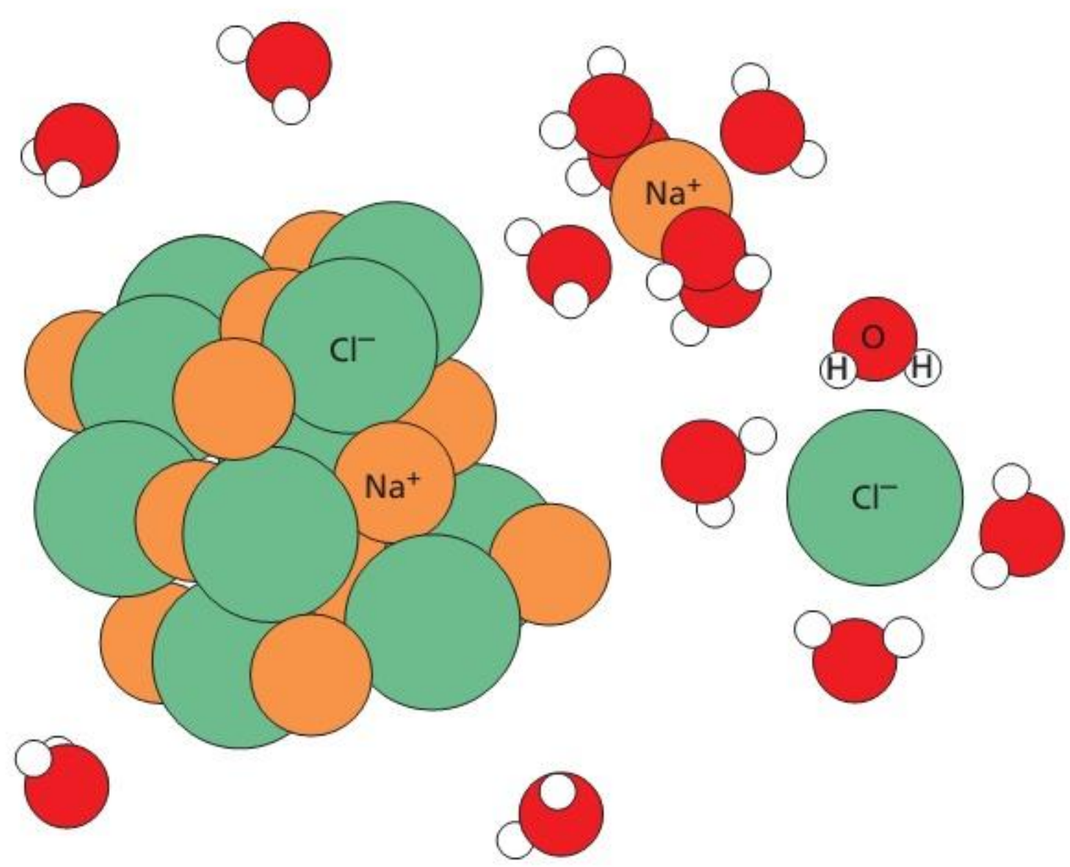


Figure 3.17 Salt (an ionic compound) dissolving in water.

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 Na^+ ion for every Cl^- ion, giving rise to the formula 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 $MgCl_2$. This formula is arrived at by each Mg^{2+} ion combining with two 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**. Na^+ has a valency of 1, but Mg^{2+} has a valency of 2. Na^+ can bond (combine) with only one Cl^- ion, whereas Mg^{2+} can bond with two Cl^- ions.

Some elements, such as copper and iron, possess two ions with different valencies. Copper can form the Cu^+ ion and the Cu^{2+} ion, with valencies 1 and 2 respectively. Therefore it can form two different compounds with chlorine, CuCl and $CuCl_2$. We can also distinguish the difference by using Roman numerals in their names: CuCl is copper(I) chloride and $CuCl_2$ is copper(II) chloride. Similarly, iron forms the Fe^{2+} and Fe^{3+} ions and so can also form two different compounds with, for example, chlorine: $FeCl_2$ (iron(II) chloride) and $FeCl_3$ (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 $Mg(NO_3)_2$. You will notice that the 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:

$$Mg\,(N\,O_3)_2$$
$$1Mg:2N:6O$$

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 I_2 but an oxidation state of −1 in I^-.

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

	Valency (oxidation state)					
	1		2		3	
Metals	Lithium	(Li^+)	Magnesium	(Mg^{2+})	Aluminium	(Al^{3+})
	Sodium	(Na^+)	Calcium	(Ca^{2+})	Iron(III)	(Fe^{3+})
	Potassium	(K^+)	Copper(II)	(Cu^{2+})		
	Silver	(Ag^+)	Zinc	(Zn^{2+})		
	Copper(I)	(Cu^+)	Iron(II)	(Fe^{2+})		
			Lead	(Pb^{2+})		
			Barium	(Ba^{2+})		
Non-metals	Fluoride	(F^-)	Oxide	(O^{2-})		
	Chloride	(Cl^-)	Sulfide	(S^{2-})		
	Bromide	(Br^-)				
	Hydrogen	(H^+)				
Groups of atoms	Hydroxide	(OH^-)	Carbonate	(CO_3^{2-})	Phosphate	(PO_4^{3-})
	Nitrate	(NO_3^-)	Sulfate	(SO_4^{2-})		
	Ammonium	(NH_4^+)	Dichromate(VI)	($Cr_2O_7^{2-}$)		
	Hydrogencarbonate	(HCO_3^-)				
	Manganate(VII)	(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 Fe^{2+} to Fe^{3+}, is oxidation.
- However, a reduction in the oxidation state, for example from +6 to +3 as in the case of Cr^{6+} (in CrO_4^{2-}) to 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 → magnesium sulfate + hydrogen

$$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$$

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.
   a $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$
   b $Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$
   c $Mg(s) + Cu(NO_3)_2(aq) \rightarrow Mg(NO_3)_2(aq) + Cu(s)$
   d $Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$

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:

$$Al = 3 \quad \text{and} \quad S = 2$$

Hence the chemical formula is Al_2S_3.

Similarly, in sodium sulfate the valencies are:

$$Na = 1 \quad \text{and} \quad SO_4 = 2$$

Hence the chemical formula is Na_2SO_4.

Questions

1 Using the information in Table 3.6, write the formulae for:
- a copper(I) oxide
- b zinc phosphate
- c iron(III) chloride
- d lead bromide
- e potassium manganate(VII)
- f sodium dichromate(VI).

2 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, 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:

H—H

A similar example exists in the diatomic halogen molecule chlorine, Cl_2 (Figure 3.21, p. 46).

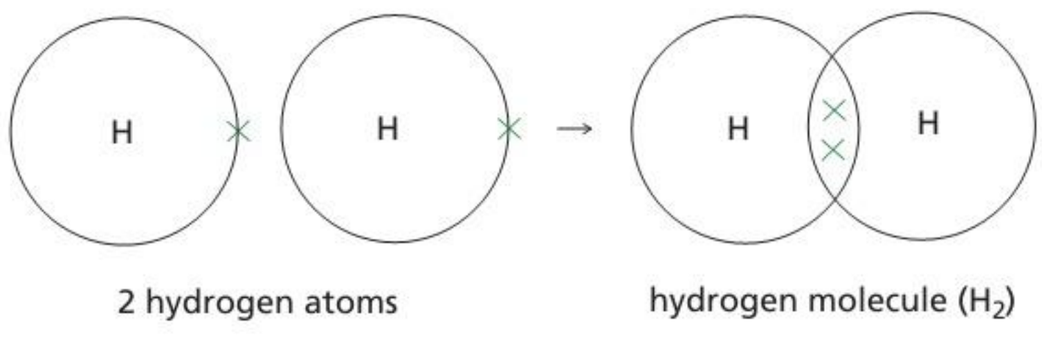


Figure 3.20a The electron sharing to form the single covalent bond in H_2 molecules.

Figure 3.20b Model of hydrogen molecule.

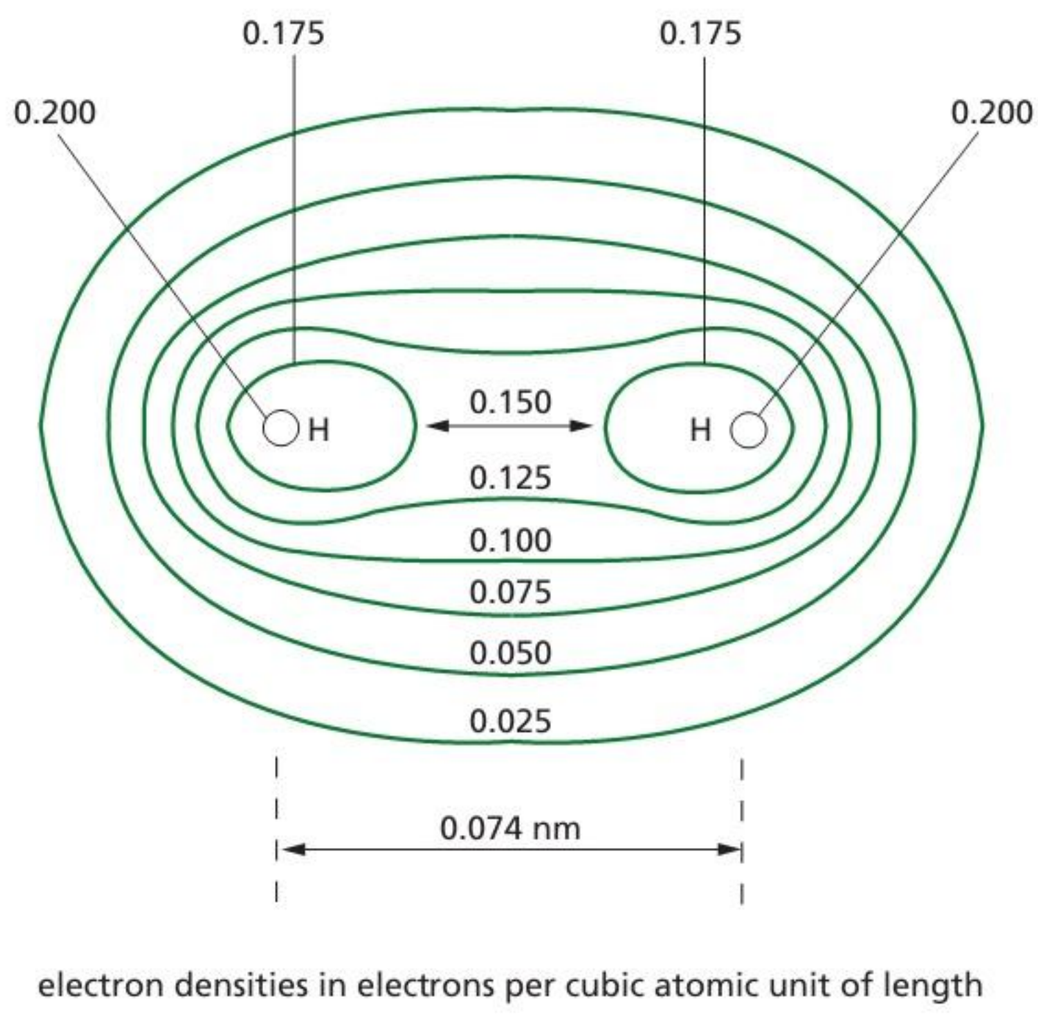


Figure 3.20c Electron density map of a hydrogen molecule.

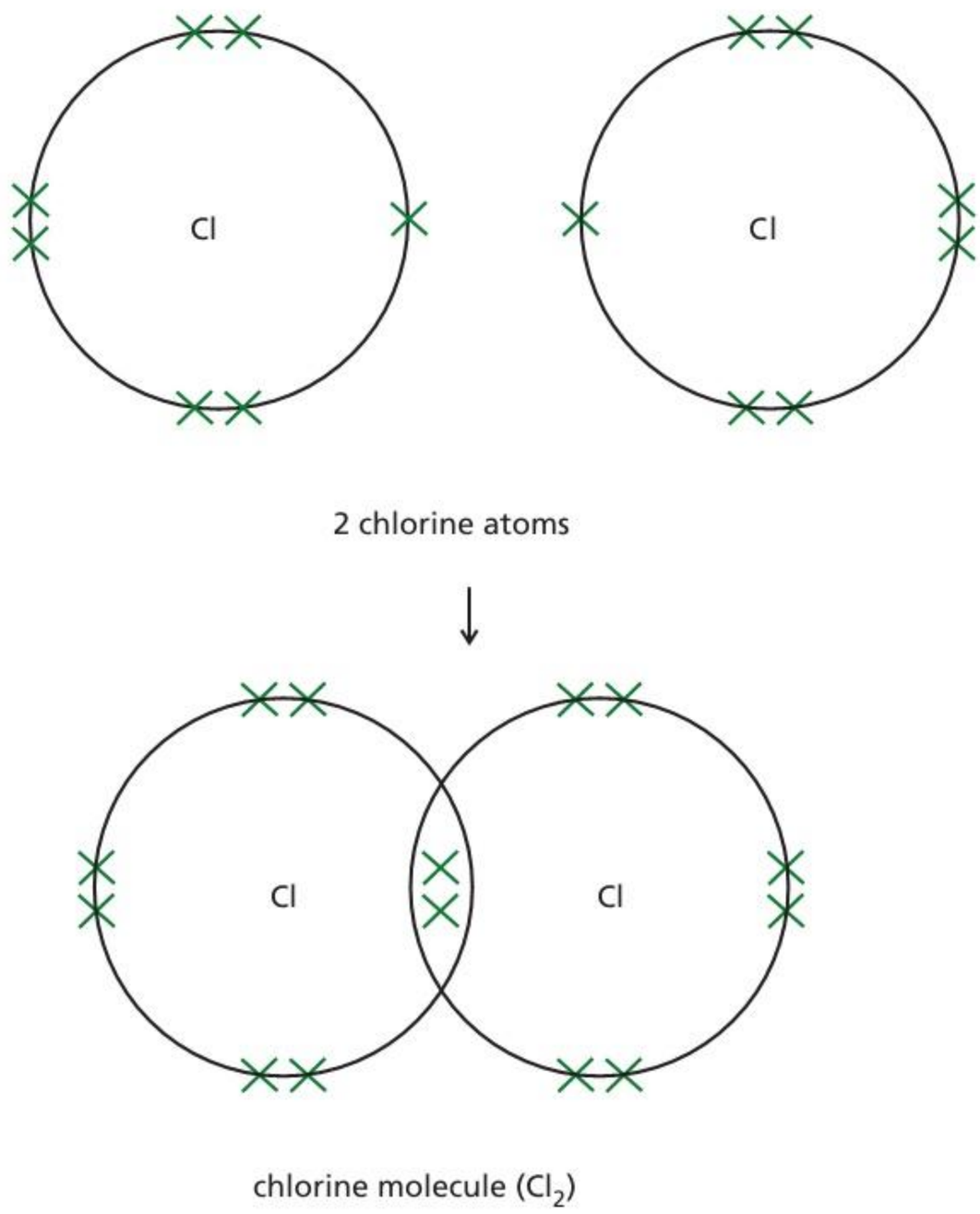


Figure 3.21 The electron sharing to form the single covalent bond in Cl_2 molecules (sharing outer electron shells only).

Other covalent compounds

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

$$_{6}C\ 2,4 \qquad _{1}H\ 1$$

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.

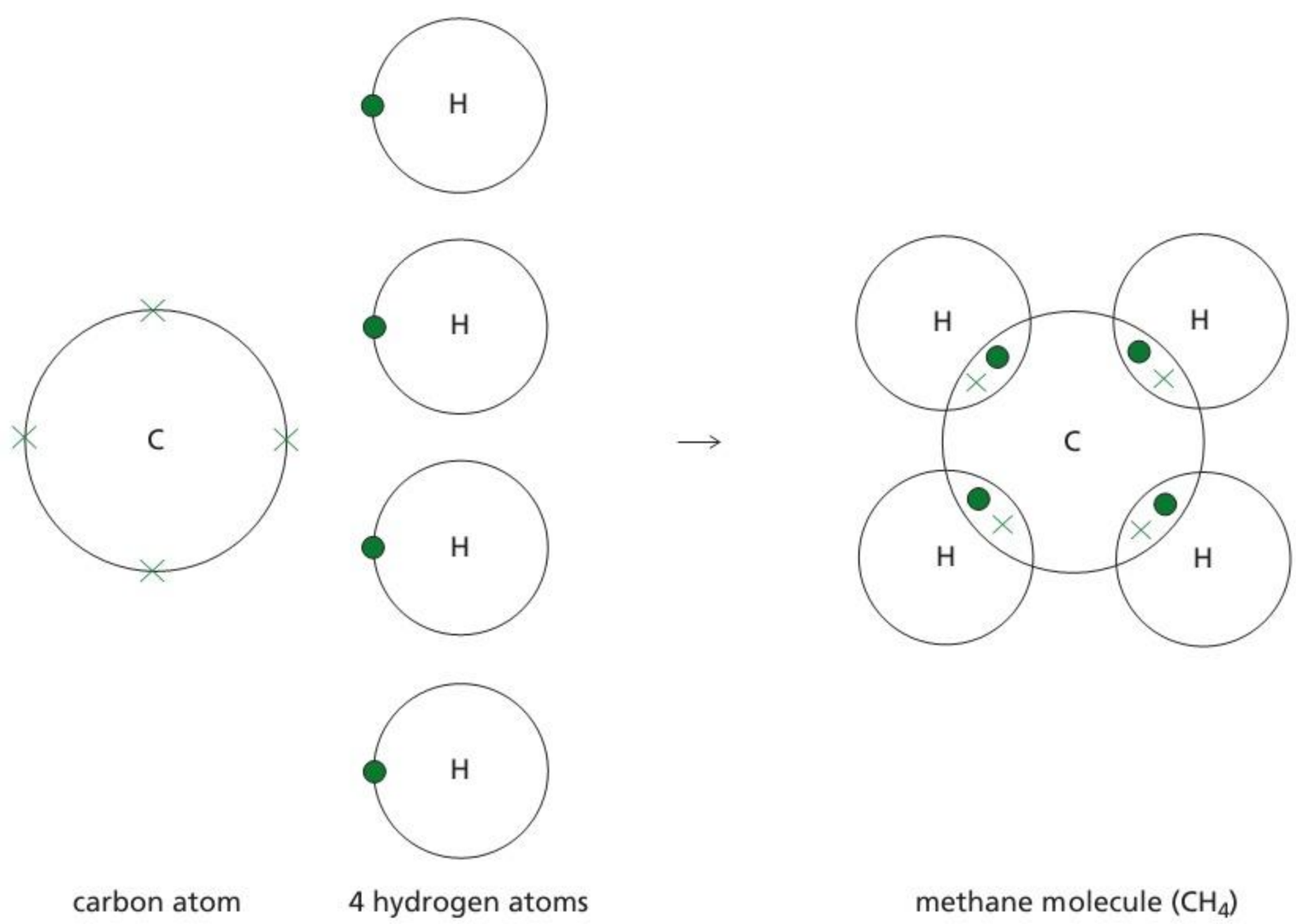
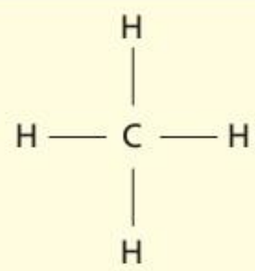
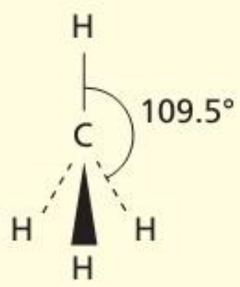


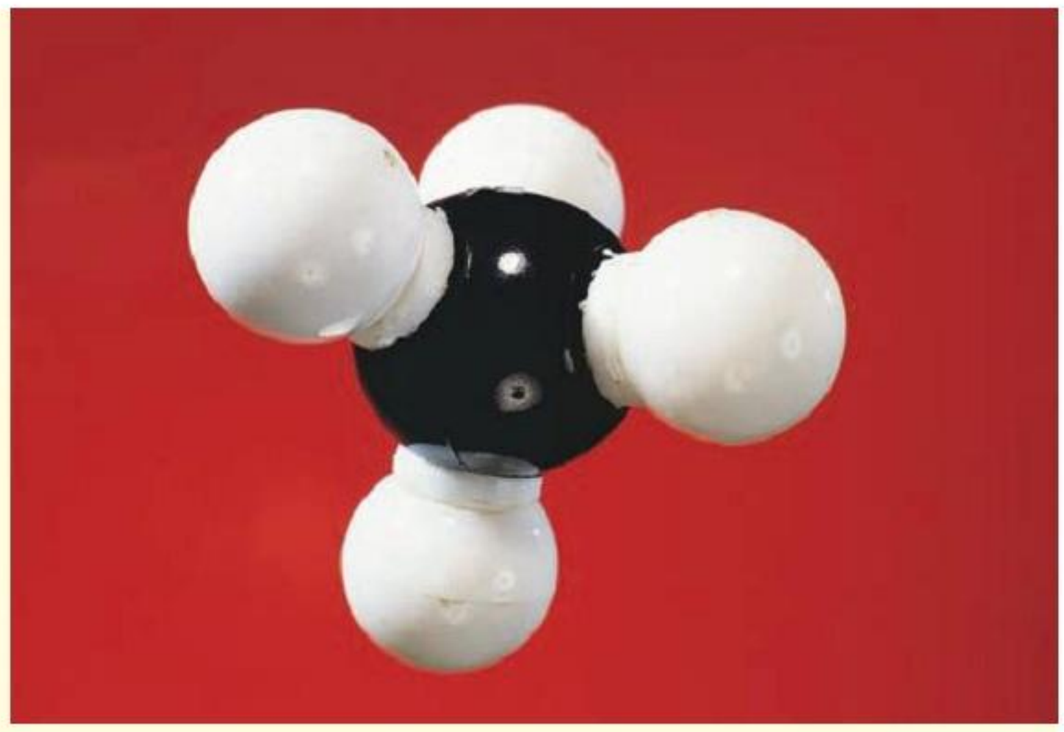
Figure 3.22 Formation of methane.



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:

$_{7}$N 2,5 $_{1}$H 1

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:

$_{8}$O 2,6 $_{1}$H 1

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).

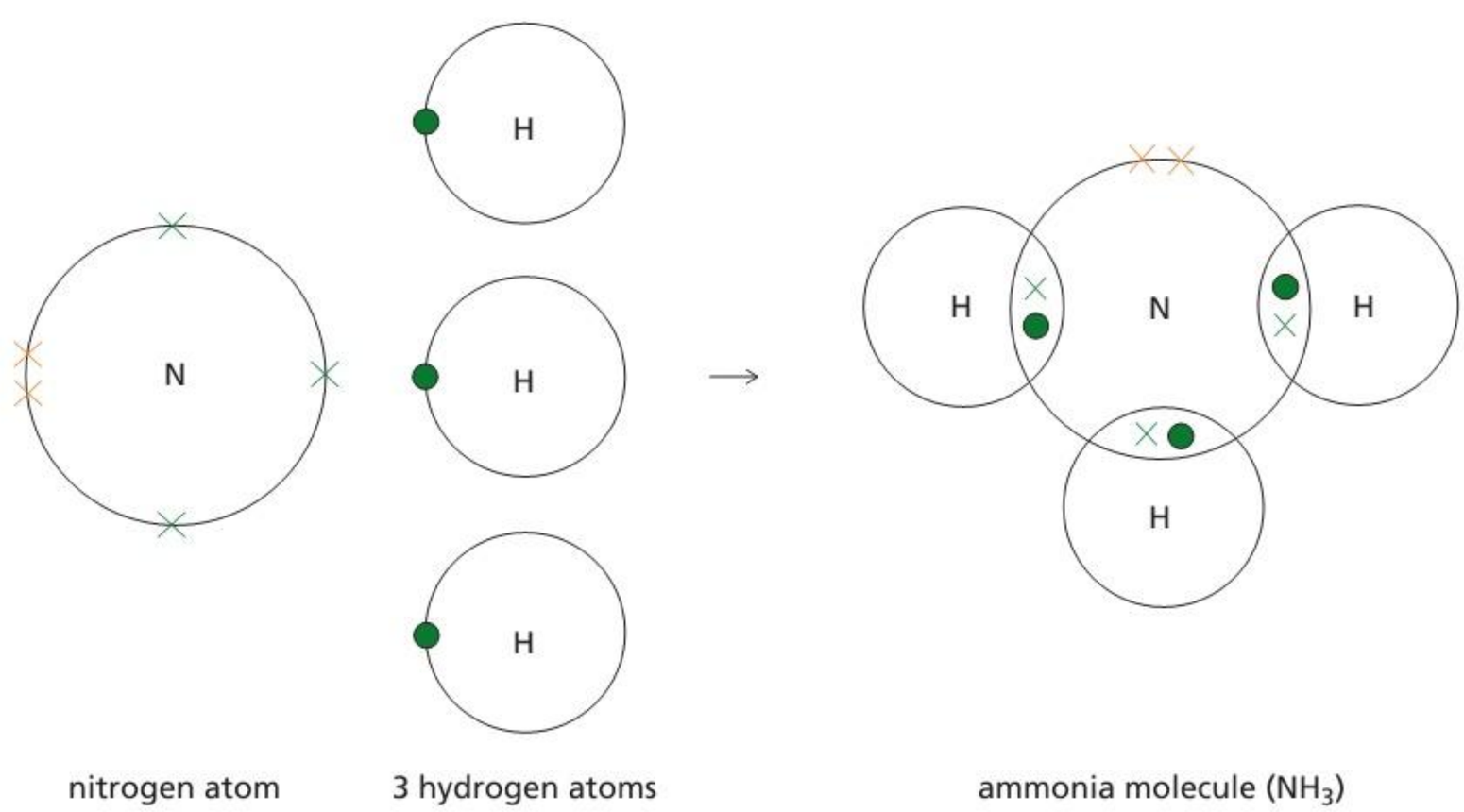
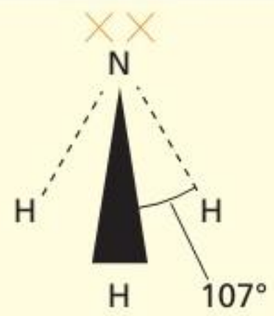
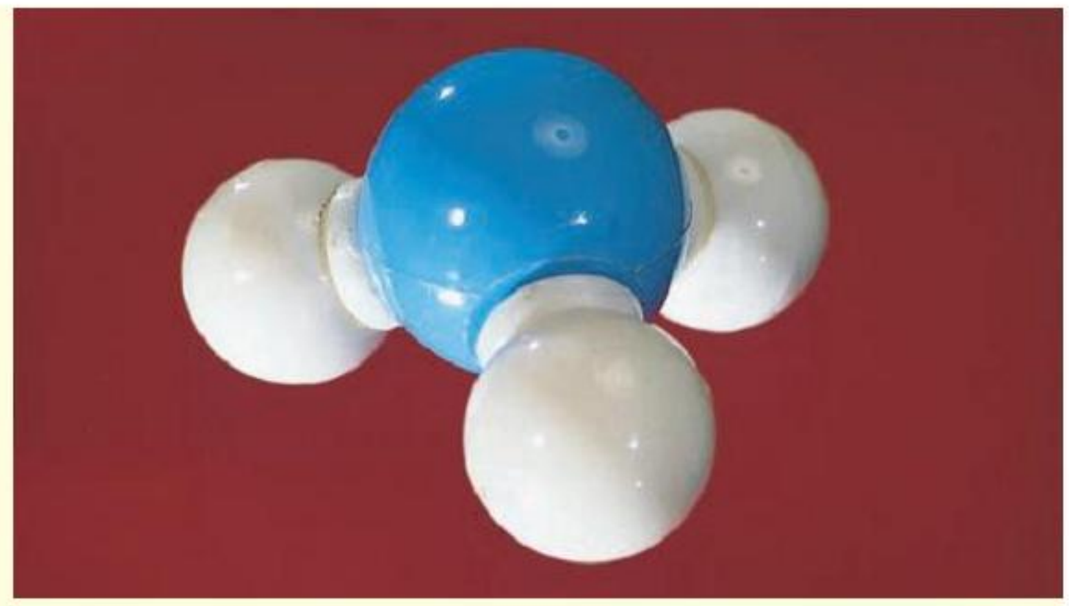


Figure 3.24 Formation of ammonia.



a Ammonia molecule.



b Model of the ammonia molecule. The pyramidal shape is caused by the repulsion between the bonding pairs of electrons as well as the lone pair (or non-bonding pair) of electrons.

Figure 3.25

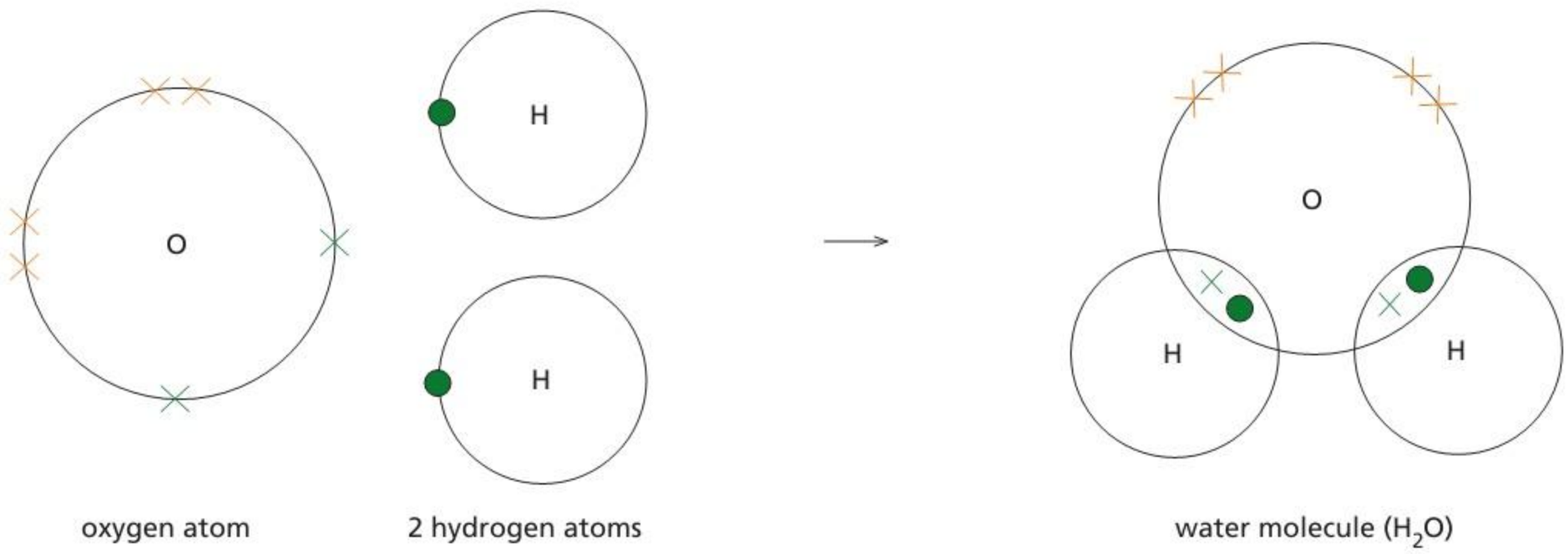
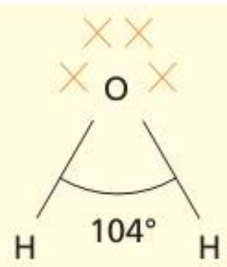
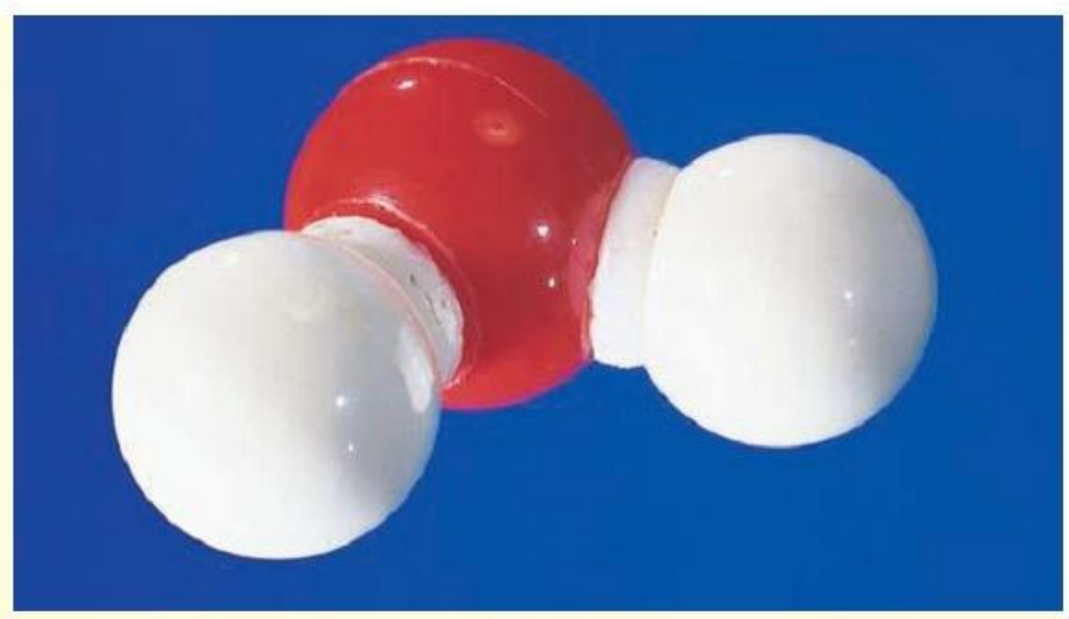


Figure 3.26 Formation of water.



a Water molecule.



b Model of a water molecule. This is a V-shaped molecule.

Figure 3.27

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

$_6C$ 2,4 $_8O$ 2,6

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).

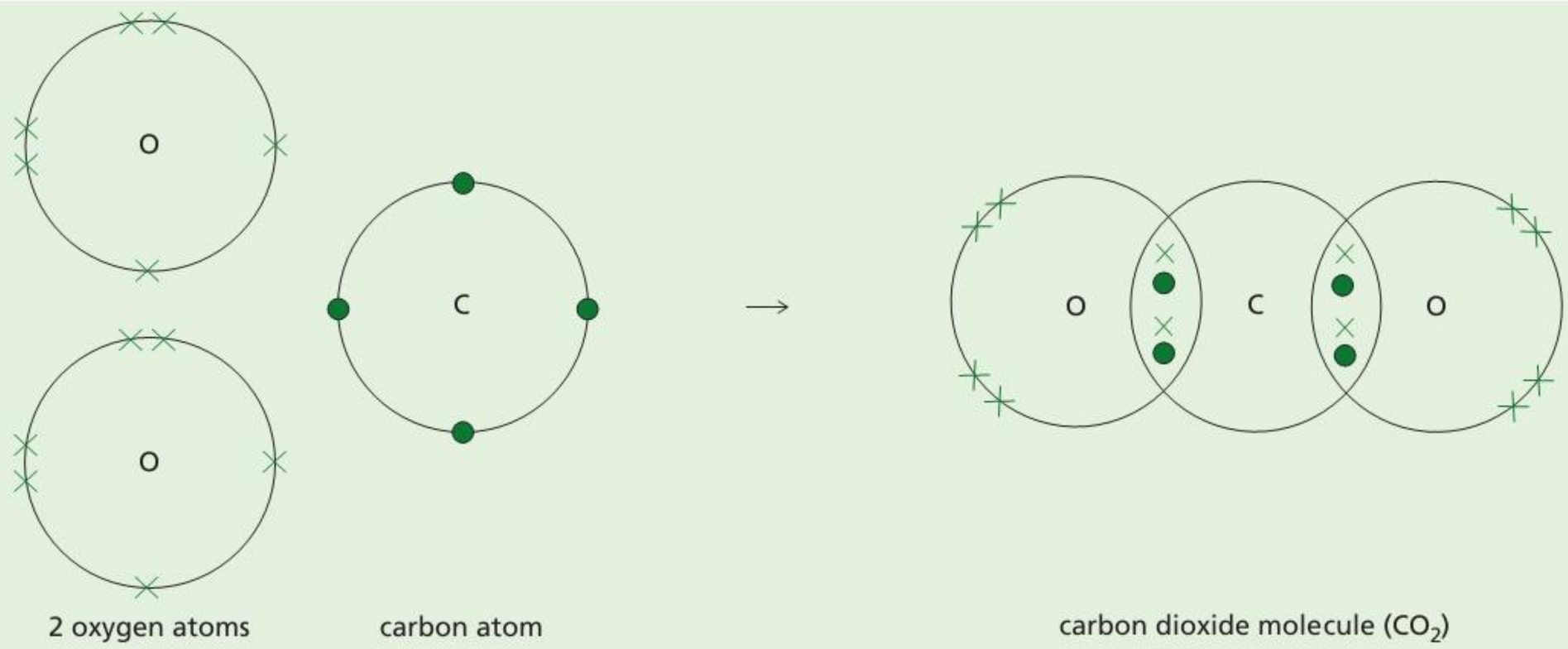
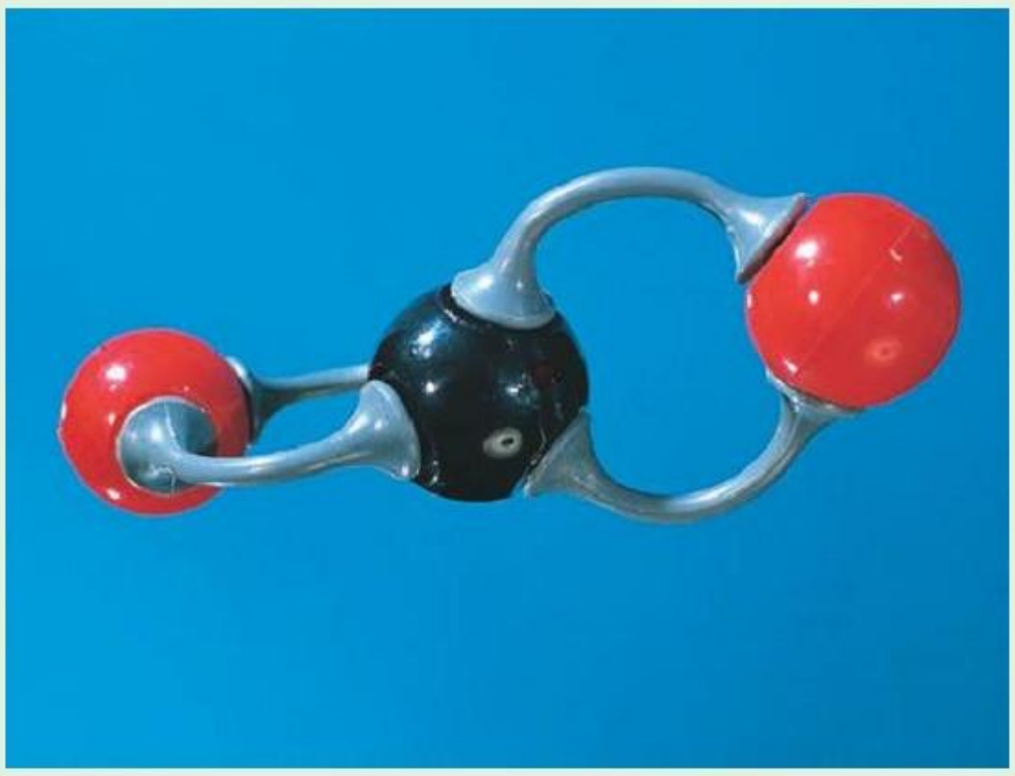


Figure 3.28 Formation of carbon dioxide.

O═C═O

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

1 Draw diagrams to represent the bonding in each of the following covalent compounds:
   - **a** tetrachloromethane (CCl_4)
   - **b** oxygen gas (O_2)
   - **c** hydrogen sulfide (H_2S)
   - **d** hydrogen chloride (HCl)
   - **e** ethene (C_2H_4)
   - **f** methanol (CH_3OH)
   - **g** nitrogen (N_2).

2 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.

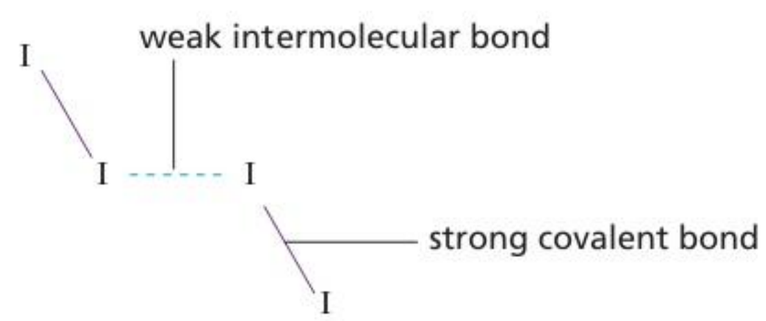
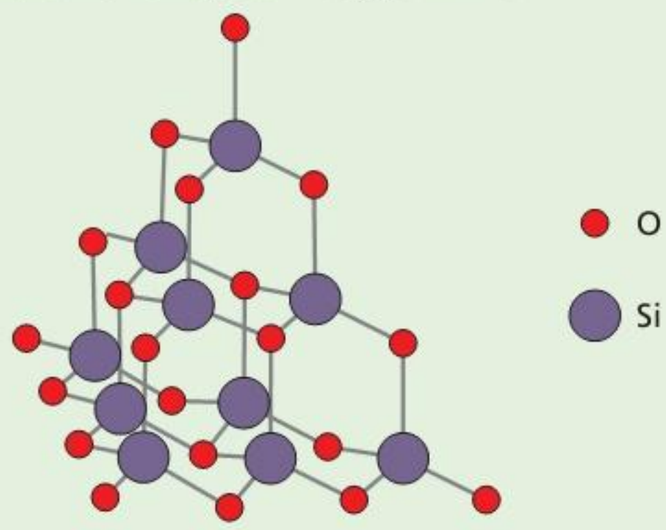


Figure 3.30 Strong covalent and weak intermolecular forces in iodine.

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:

$$HCl(g) \xrightarrow{\text{water}} H^+(aq) + Cl^-(aq)$$

- 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^{-3}	2.25	3.51

Table 3.8 Uses of graphite and diamond.

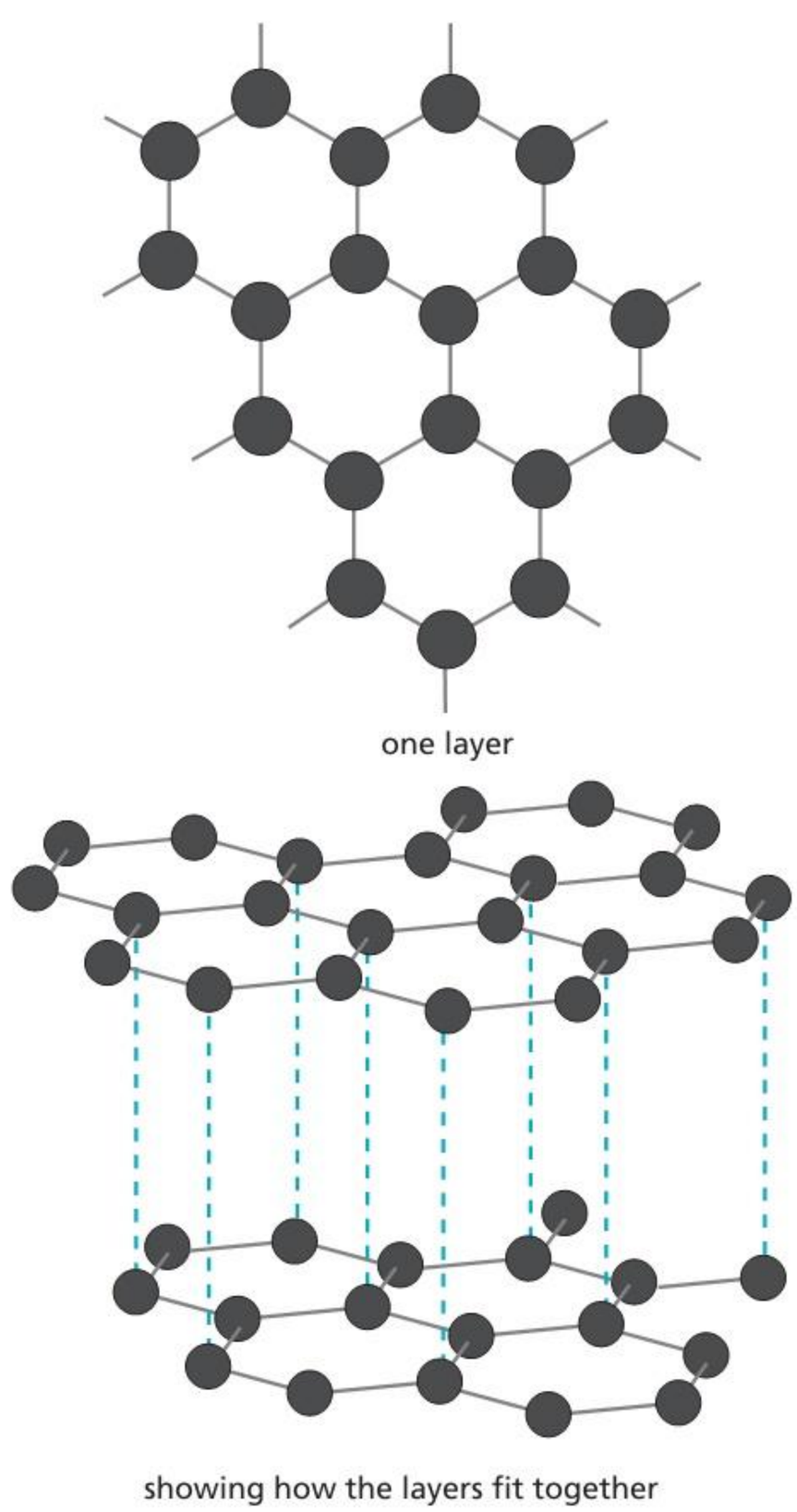
Graphite	Diamond
Pencils	Jewellery
Electrodes	Glass cutters
Lubricant	Diamond-studded saws
	Drill bits
	Polishers

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 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.



a A portion of the graphite structure.

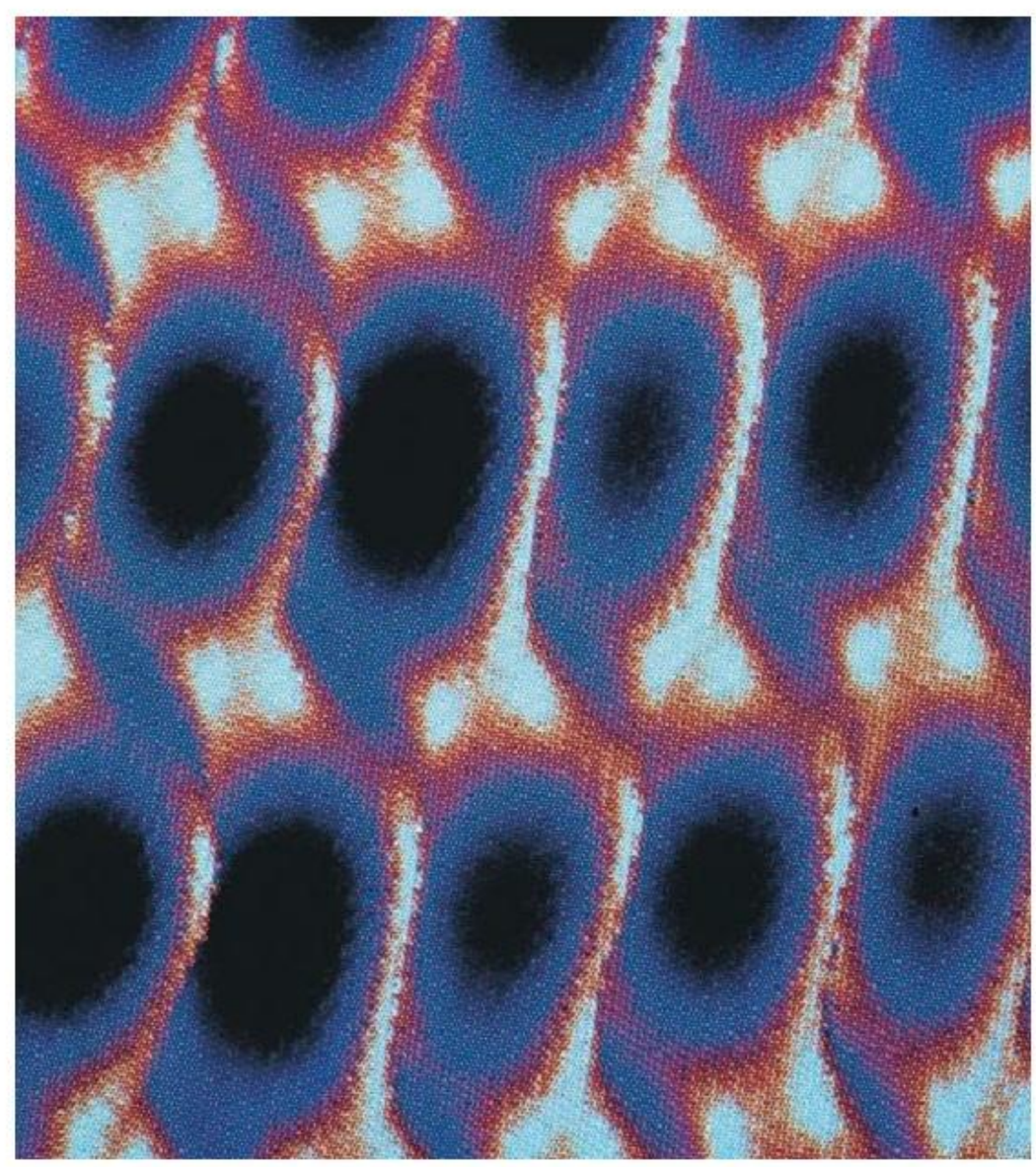
Figure 3.32

Graphene

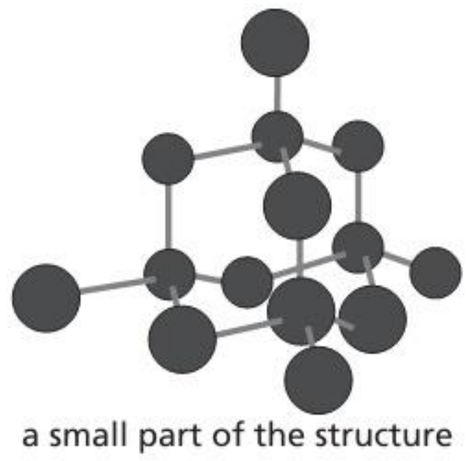
Discovered in 2004, graphene is a so-called super material made up of single layers of graphite as shown in the upper diagram of Figure 3.32a. It is able to conduct electricity one million times better than copper metal and has enormous potential in electronics.

Diamond

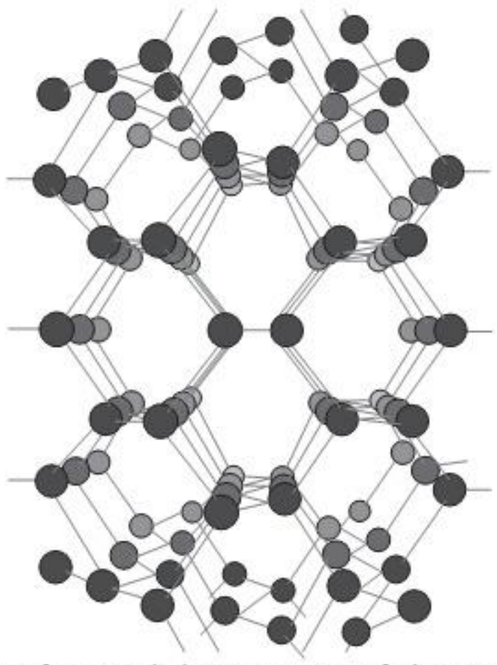
Figure 3.33a shows the diamond structure. Each of the carbon atoms in the giant structure is covalently bonded to four others. They form a tetrahedral arrangement similar to that found in silicon(IV) oxide (p. 50). This bonding scheme gives rise to a very



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



a small part of the structure



a view of a much larger part of the structure

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).

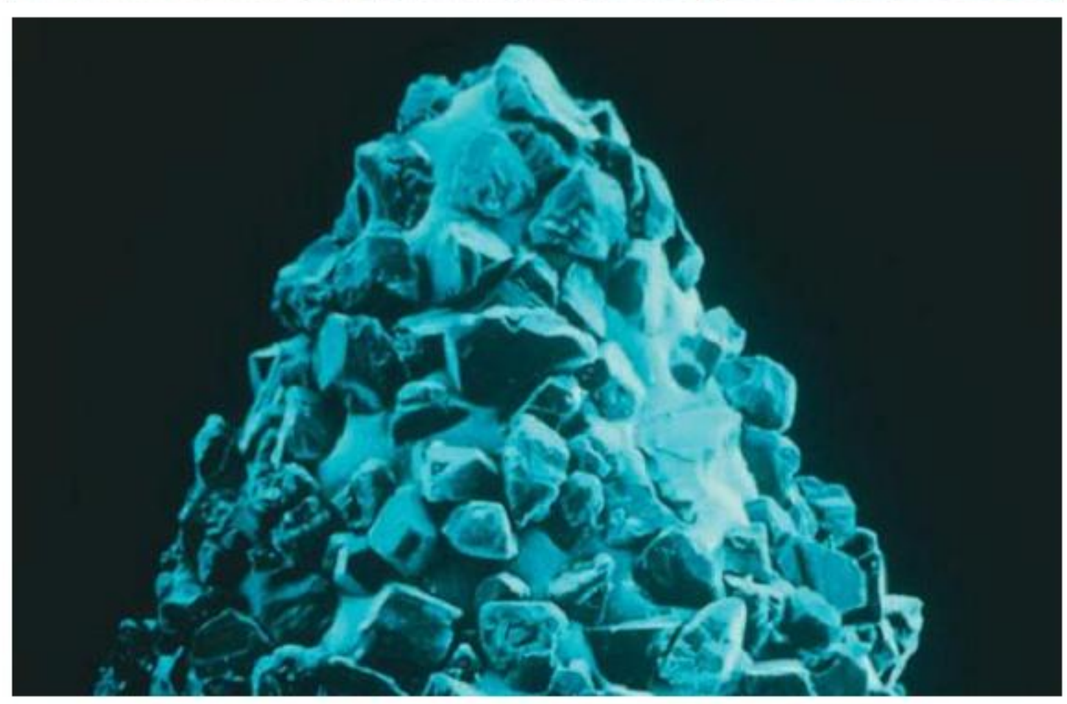


Figure 3.34 Uses of graphite (as a pencil 'lead' and in a squash racket) and diamond (as a toothed saw to cut marble and on a dentist's drill).

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.

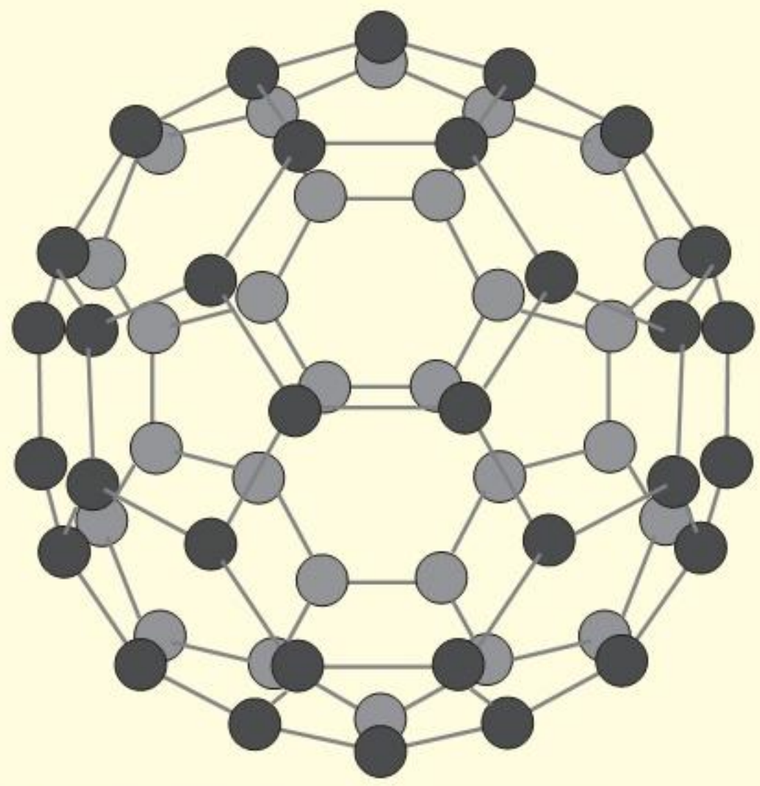


Figure 3.35 Buckminsterfullerene – a 'bucky ball' (C_{60}).

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**.

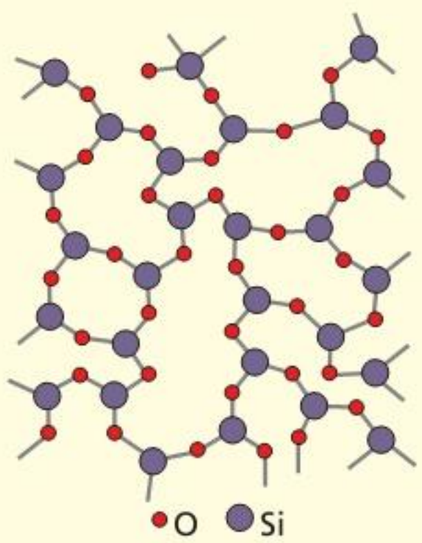


Figure 3.37 Two-dimensional structure of silicon(IV) oxide.

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, $Al_2Si_2O_5(OH)_4$, 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 °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 (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).

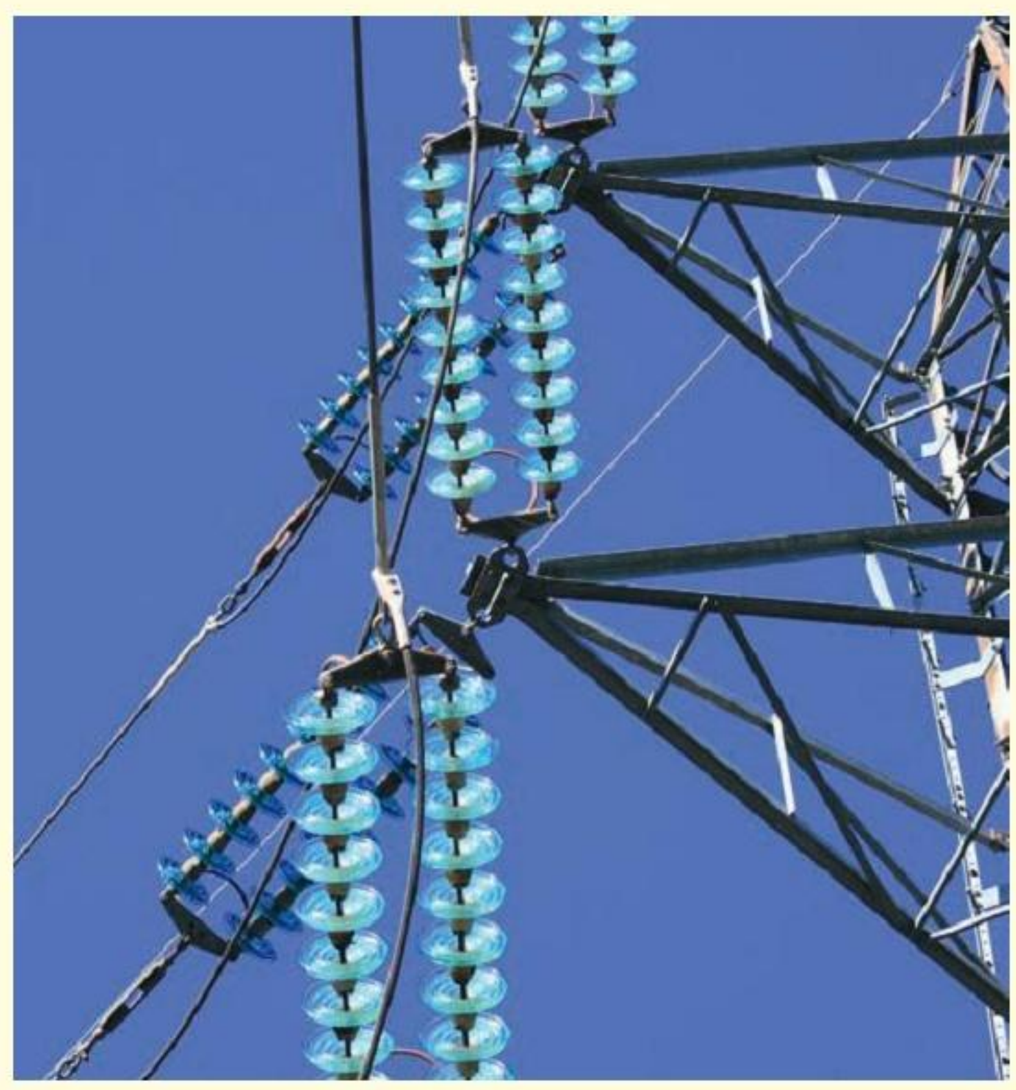


Figure 3.39 Ceramic discs prevent short-circuits in the national grid.

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**.

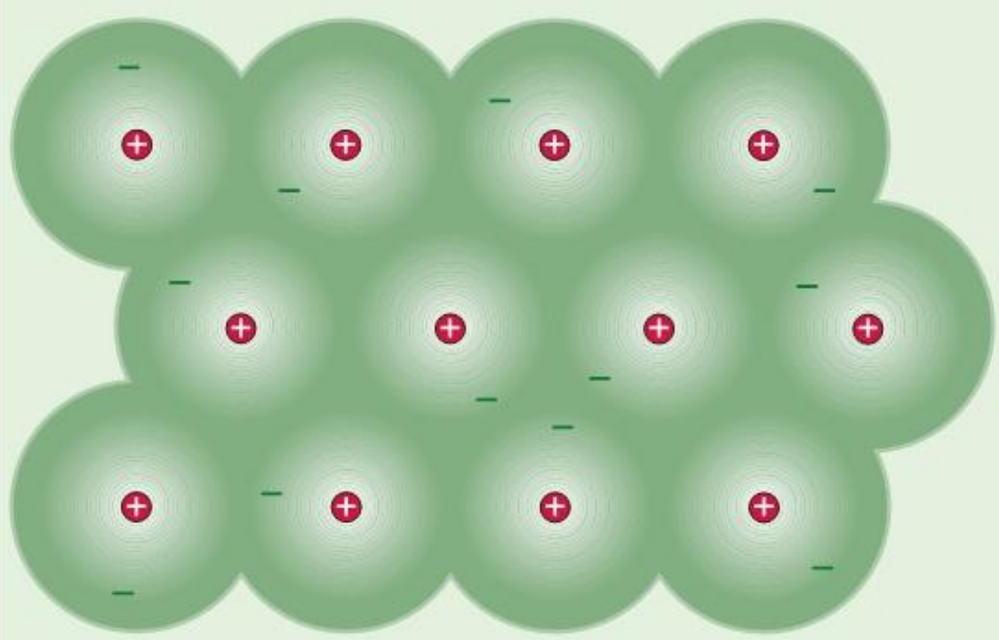


Figure 3.40 Metals consist of positive ions surrounded by a 'sea' of electrons.

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.

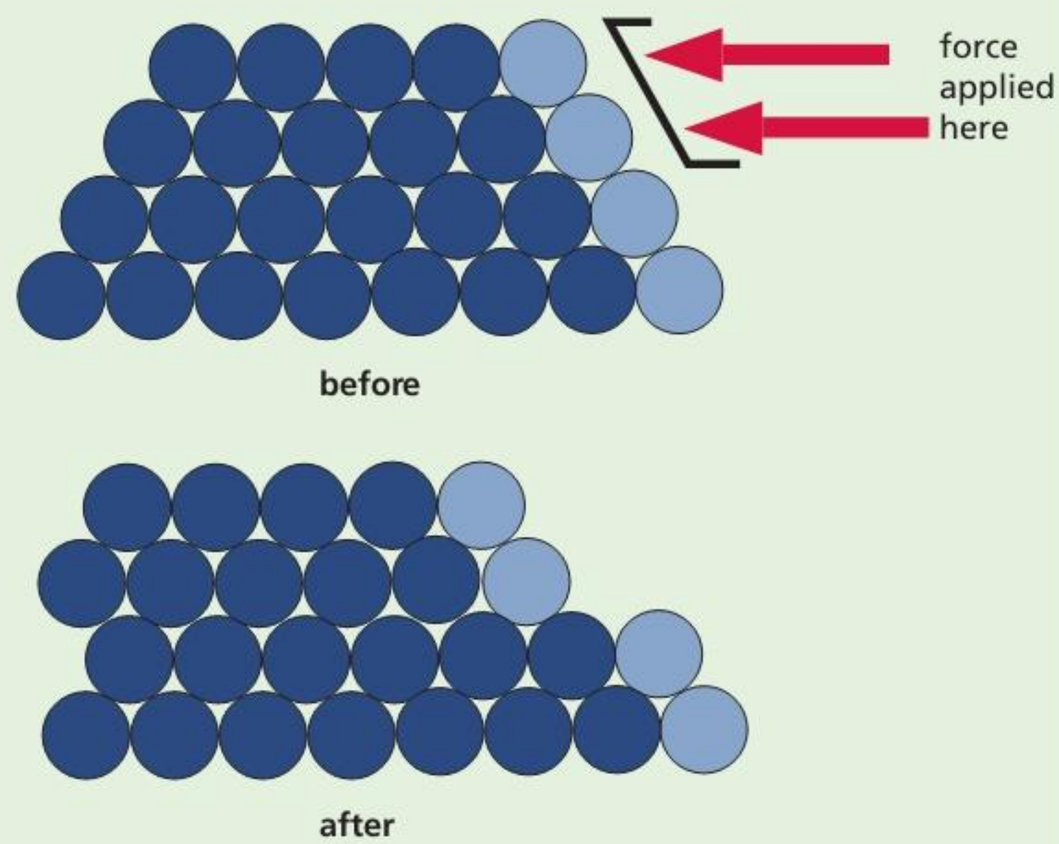


Figure 3.41 The positions of the positive ions in a metal before and after a force has been applied.

- 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.

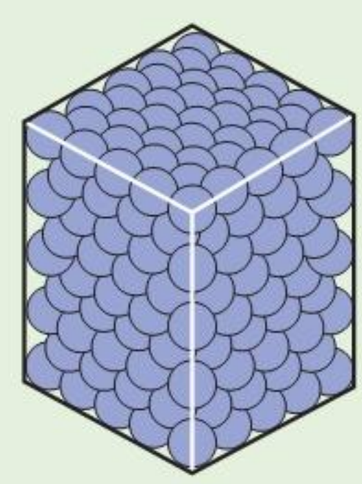
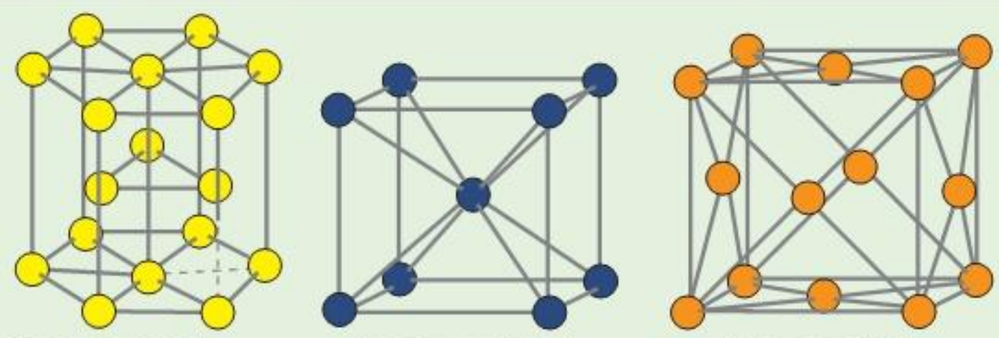


Figure 3.42 Arrangement of ions in the crystal lattice of a metal.



close-packed hexagonal structure (e.g. magnesium, density = 1.74 g cm^{-3})

body-centred cubic structure (e.g. iron, density = 7.87 g cm^{-3})

face-centred cubic structure (e.g. copper, density = 8.92 g cm^{-3})

Figure 3.43 Relating different structures to the density of metals.

Questions

1 Explain the terms:
   a malleable b ductile.
2 Explain why metals are able to conduct heat and electricity.
3 Explain why the melting point of magnesium (649 °C) is much higher than the melting point of sodium (97.9 °C).
4 Suggest reasons why plastics and ceramics are used to manufacture handles for metallic cooking utensils.

Checklist

After studying Chapter 3 you should know and understand the following terms.

- **Allotropy** The existence of an element in two or more different forms in the same physical state.
- **Atomic mass unit** Exactly $\frac{1}{12}$ of the mass of one atom of the most abundant isotope of carbon-12.
- **Ceramics** Materials such as pottery made from inorganic chemicals by high-temperature processing. Other modern ceramics include zirconium oxide and silicon nitride.
- **Covalent bond** A chemical bond formed by the sharing of one or more pairs of electrons between two atoms.
- **Delocalised** Refers to spreading out of electrons within the metal structure.
- **Electron** A fundamental sub-atomic particle with a negative charge present in all atoms within energy levels around the nucleus.
- **Electron energy levels (shells)** The allowed energies of electrons in atoms.
- **Electronic structure (configuration)** A shorthand method of describing the arrangement of electrons within the energy levels of an atom.
- **Electrostatic force of attraction** A strong force of attraction between opposite charges.
- **Giant ionic structure** A lattice held together by the electrostatic forces of attraction between ions.
- **Giant molecular or macromolecular substance** A substance containing thousands of atoms per molecule.
- **Glass** A supercooled liquid which forms a hard, brittle substance that is usually transparent and resistant to chemical attack.
- **Intermolecular bonds** Attractive forces which act between molecules, for example van der Waals' forces.
- **Intramolecular bonds** Forces which act within a molecule, for example covalent bonds.
- **Ion** An atom or group of atoms which has either lost one or more electrons, making it positively charged, or gained one or more electrons, making it negatively charged.
- **Ionic (electrovalent) bond** A strong electrostatic force of attraction between oppositely charged ions.
- **Ionisation** The process whereby an atom gains or loses an electron(s) to become an ion.
- **Isotopes** Atoms of the same element which possess different numbers of neutrons. They differ in nucleon number (mass number).
- **Lattice** A regular three-dimensional arrangement of atoms/ions in a crystalline solid.
- **Mass spectrometer** A device in which atoms or molecules are ionised and then accelerated. Ions are separated according to their mass.
- **Metallic bond** An electrostatic force of attraction between the mobile 'sea' of electrons and the regular array of positive metal ions within the solid metal.
- **Neutron** A fundamental, uncharged sub-atomic particle present in the nuclei of atoms.
- **Nucleon number (mass number)** Symbol *A*. The total number of protons and neutrons found in the nucleus of an atom.
- **Nucleus** Found at the centre of the atom, it contains the protons and neutrons.
- **Oxidising agent** A substance that causes an increase in oxidation number.
- **Oxidation** The process of removing an electron or electrons from an atom to create a positive ion. Gives an increase in oxidation number.
- **Proton** A fundamental sub-atomic particle which has a positive charge equal in magnitude to that of an electron. Protons occur in all nuclei.
- **Proton number (atomic number)** Symbol *Z*. The number of protons in the nucleus of an atom. The number of electrons present in an atom. The order of the element within the Periodic Table (p. 136).
- **Radioactive** A property of unstable isotopes. They disintegrate spontaneously to give off one or more types of radiation.
- **Radioisotope** A radioactive isotope.
- **Reducing agent** A substance that causes a decrease in oxidation number.
- **Reduction** The process of adding an electron or electrons to an atom to create a negative ion. Gives a decrease in oxidation number.

- **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 $^{69}_{31}Ga$ and $^{71}_{31}Ga$ 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 $^{69}_{31}Ga$ and 40% of atoms of $^{71}_{31}Ga$. 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 (CaF_2)

b oxygen (O_2)

c magnesium chloride ($MgCl_2$)

d tetrachloromethane (CCl_4).

4 Use the information given in Table 3.6 on p. 43 to work out the formula for:

a silver oxide

b zinc chloride

c potassium sulfate

d calcium nitrate

e iron(II) nitrate

f copper(II) carbonate

g iron(III) hydroxide

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**

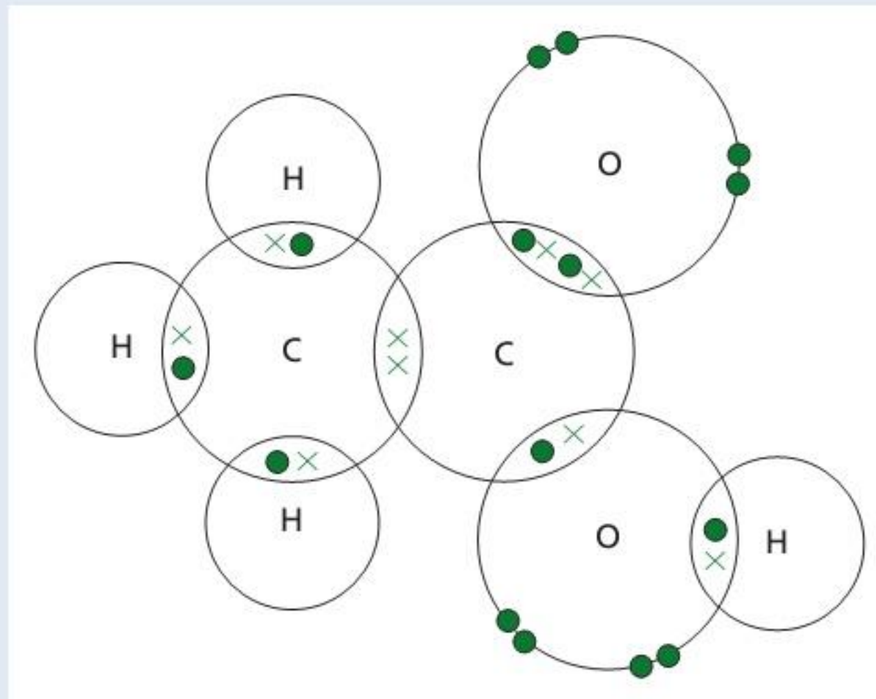
(ii) **Y** and **Z**

(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.

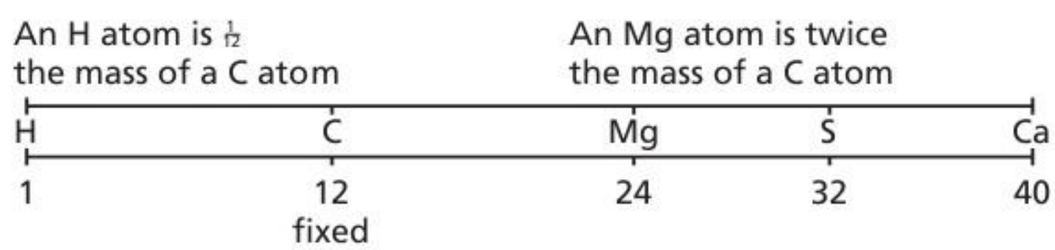


Figure 4.1 The relative atomic masses of the elements H, C, Mg, S and Ca.

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?

$$C + O_2 \rightarrow CO_2$$

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 O_2 and 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.

O_2 has a relative formula mass of $2 \times 16 = 32$
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.

C	+	O_2	→	CO_2
12		32		44

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

C	+	O_2	→	CO_2
12		32		44
12 g		32 g		44 g

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:

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

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, H_2O, is
$(2 \times 1) + 16 = 18$
Relative formula mass of sodium hydroxide is
$23 + 16 + 1 = 40$
Relative formula mass of hydrogen gas, 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.

2Na	+	$2H_2O$	→	2NaOH	+	H_2
$2 \times 23 = 46$		$2 \times 18 = 36$		$2 \times 40 = 80$		2

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

$2Na$	+	$2H_2O$	→	$2NaOH$	+	H_2
$2 \times 23 = 46$		$2 \times 18 = 36$		$2 \times 40 = 80$		2
46 g		36 g		80 g		2 g

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?

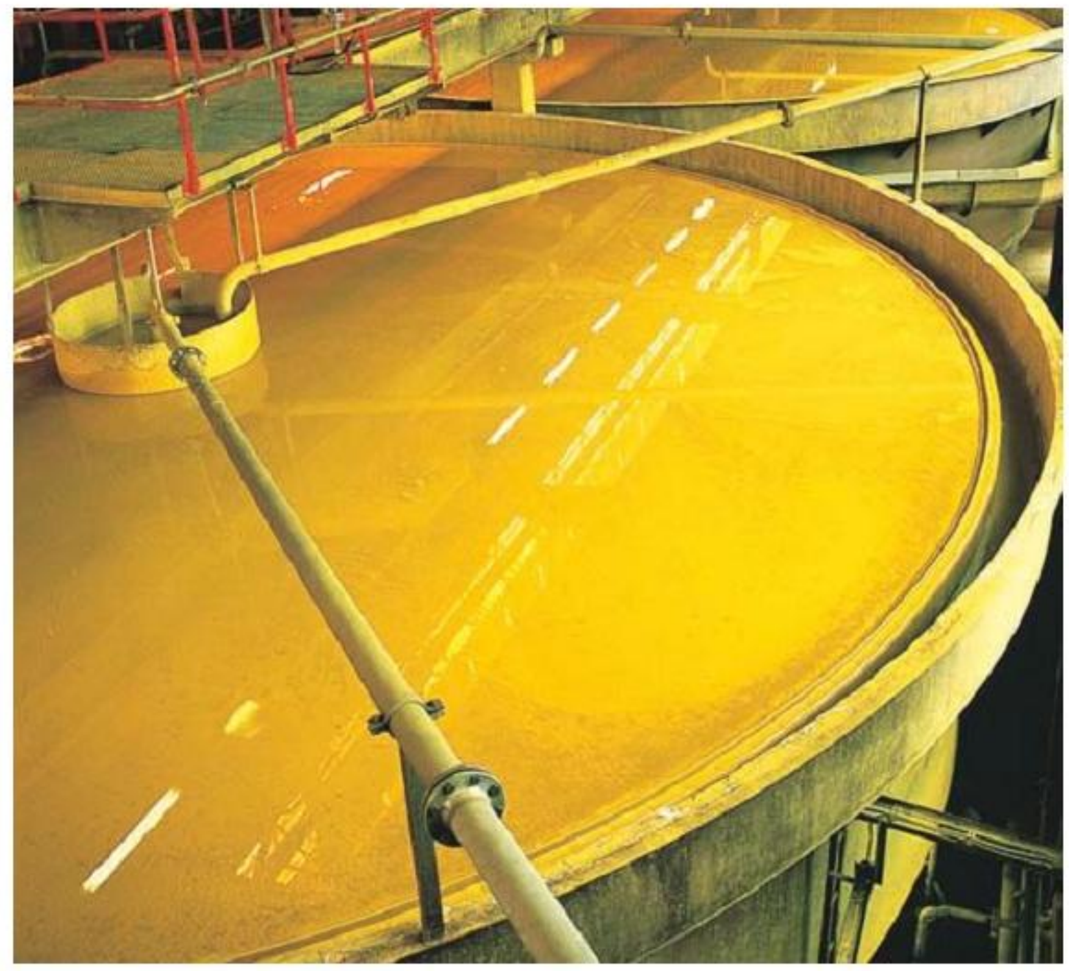


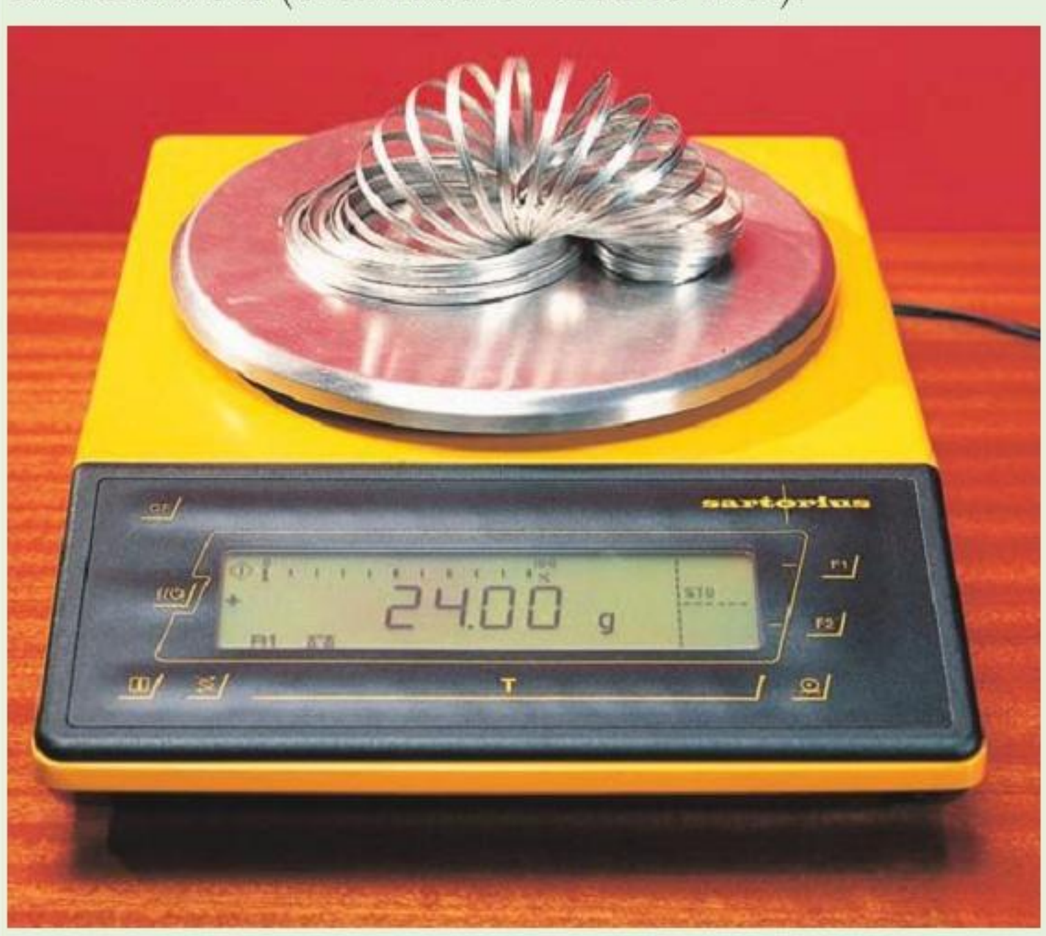
Figure 4.2 The chemists at paint manufacturers need to know how much pigment is going to be produced.

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!

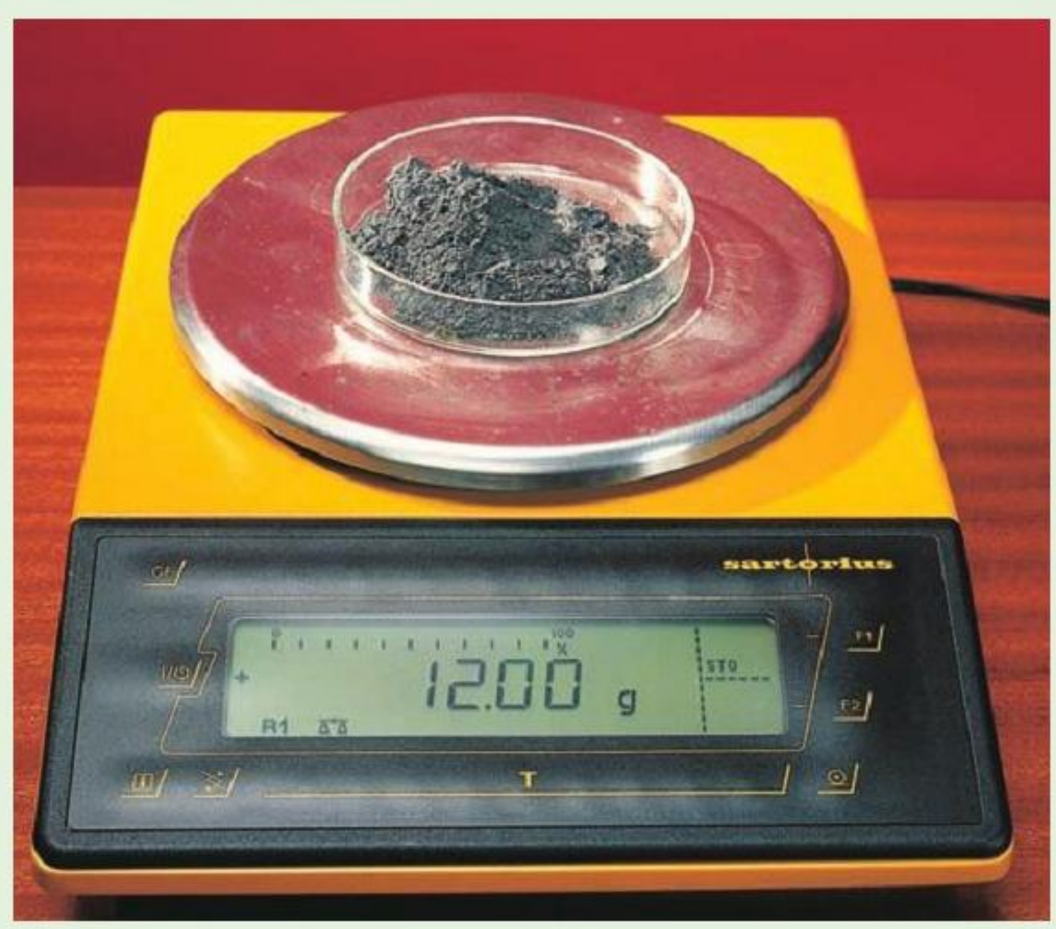
To solve this problem they need a way of counting atoms, ions or molecules. Atoms, ions and molecules are very tiny particles and it is impossible to measure out a dozen or even a hundred of them. Instead, chemists weigh out a very large number of particles.

This number is 6×10^{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 6×10^{23} particles is called a **mole** (often abbreviated to mol).

So, a mole of the element magnesium is 6×10^{23} atoms of magnesium and a mole of the element carbon is 6×10^{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 6×10^{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 6×10^{23} atoms.

The A_r for aluminium is 27. In 27 g of aluminium it is found that there are 6×10^{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
= number of moles × relative atomic mass (A_r)
$= 2 \times 56$
= 112 g

b mass of 0.25 mole of iron
= number of moles × relative atomic mass (A_r)
$= 0.25 \times 56$
= 14 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 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 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).

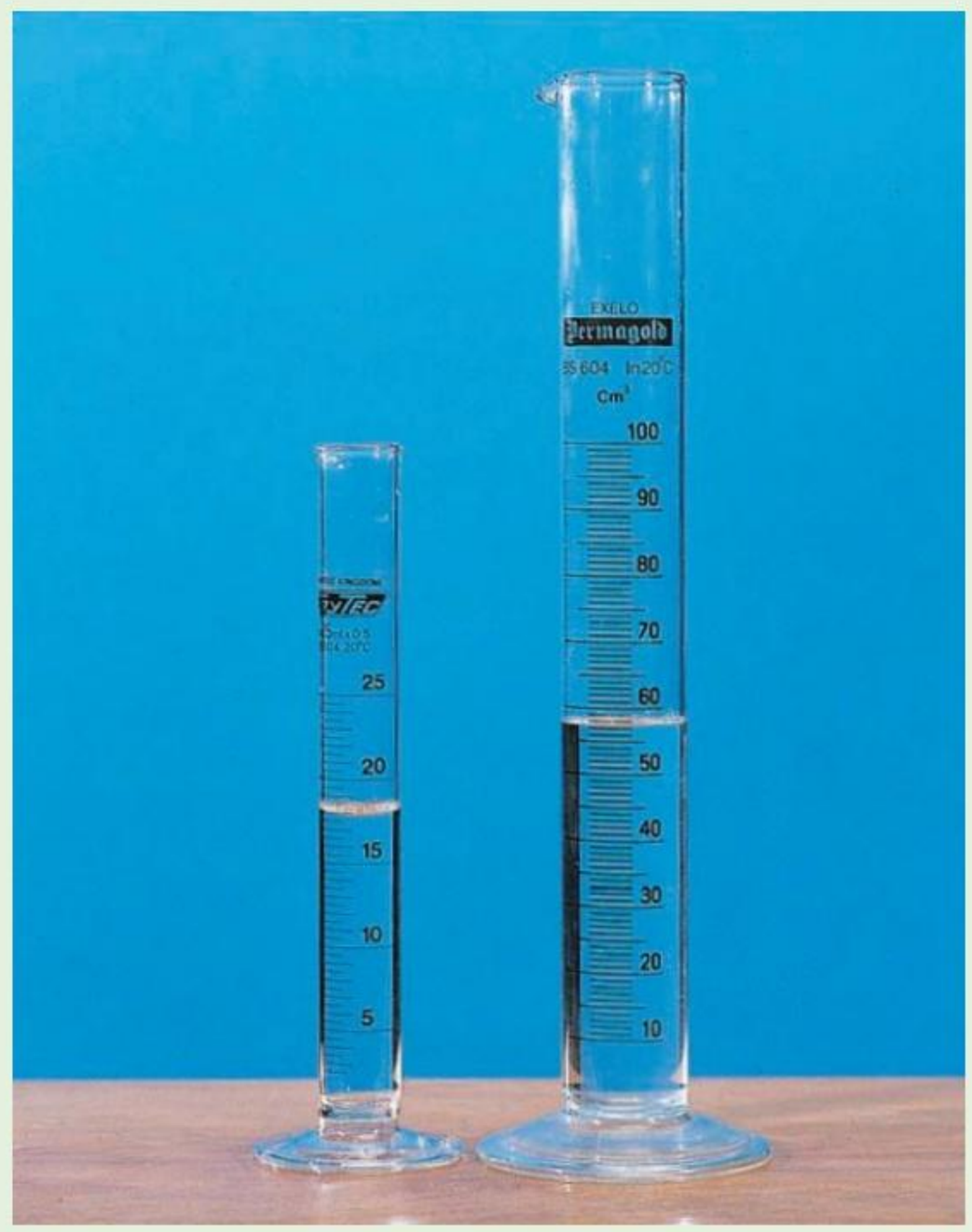


Figure 4.4 One mole of water (H_2O) (left) and one mole of ethanol (C_2H_5OH) (right) in separate measuring cylinders.

What is the mass of 1 mole of water (H_2O) molecules? (A_r: H = 1; O = 16)

From the formula of water, H_2O, 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:

$$(2 \times 1) + (1 \times 16) = 18\,g$$

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, C_2H_5OH? (A_r: H = 1; C = 12; O = 16)

a One mole of C_2H_5OH contains 2 moles of carbon atoms, 6 moles of hydrogen atoms and 1 mole of oxygen atoms. Therefore:

mass of 1 mole of ethanol
$= (2 \times 12) + (6 \times 1) + (1 \times 16)$
$= 46\,g$

b The RFM of ethanol is 46.

Example 2

What is **a** the mass of 1 mole and **b** the RFM of nitrogen gas, N_2? (A_r: N = 14)

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

mass of 1 mole of N_2
$= 2 \times 14$
$= 28\,g$

b The RFM of N_2 is 28.

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

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

Example 3

Calculate the mass of **a** 3 moles and **b** 0.2 moles of carbon dioxide gas, CO_2. (A_r: C = 12; O = 16)

a One mole of CO_2 contains 1 mole of carbon atoms and 2 moles of oxygen atoms. Therefore:

mass of 1 mole of CO_2
$= (1 \times 12) + (2 \times 16)$
$= 44\,g$

mass of 3 moles of CO_2
= number of moles × mass of 1 mole of CO_2
$= 3 \times 44$
$= 132\,g$

b mass of 0.2 mole of CO_2
= number of moles × mass of 1 mole of CO_2
$= 0.2 \times 44$
$= 8.8\,g$

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

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

Example 4

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

a One mole of MgO contains 1 mole of magnesium atoms and 1 mole of oxygen atoms. Therefore:

mass of 1 mole of MgO
$= (1 \times 24) + (1 \times 16)$
$= 40\,g$

number of moles of MgO in 80 g

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

$$= \frac{80}{40}$$

= 2 moles

b number of moles of MgO in 10 g

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

$$= \frac{10}{40}$$

= 0.25 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 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:

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

or

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

Example 1

Calculate the number of moles of ammonia gas, NH_3, in a volume of 72 dm^3 of the gas measured at rtp.

$$\text{number of moles of ammonia} = \frac{\text{volume of ammonia in dm}^3}{24\,\text{dm}^3} = \frac{72}{24} = 3$$

Example 2

Calculate the volume of carbon dioxide gas, CO_2, occupied by **a** 5 moles and **b** 0.5 mole of the gas measured at rtp.

a volume of CO_2
= number of moles of CO_2 × 24 dm^3
= 5 × 24
= 120 dm^3

b volume of CO_2
= number of moles of CO_2 × 24 dm^3
= 0.5 × 24
= 12 dm^3

The volume occupied by one mole of any gas must contain 6×10^{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 ($g\,dm^{-3}$) but more often concentration is measured in **moles per cubic decimetre ($mol\,dm^{-3}$)**. When 1 mole of a substance is dissolved in water and the solution is made up to 1 dm^3 (1000 cm^3), a **1 molar** (**1 $mol\,dm^{-3}$**) 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:

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

Example 1

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

1 mole of NaOH contains 1 mole of sodium, 1 mole of oxygen and 1 mole of hydrogen. Therefore:

mass of 1 mole of NaOH
= (1 × 23) + (1 × 16) + (1 × 1)
= 40 g

number of moles of NaOH in 10 g

$$= \frac{\text{mass of NaOH}}{\text{mass of 1 mole of NaOH}} = \frac{10}{40} = 0.25$$

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

concentration of the NaOH solution

$$= \frac{\text{number of moles of NaOH}}{\text{volume of solution (dm}^3)} = \frac{0.25}{0.25} = 1\,\text{mol dm}^{-3}$$

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:

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

Example 2

Calculate the mass of potassium hydroxide, KOH, that needs to be used to prepare $500\,cm^3$ of a $2\,mol\,dm^{-3}$ solution in water. (A_r: H = 1; O = 16; K = 39)

number of moles of KOH
= concentration of solution ($mol\,dm^{-3}$) × volume of solution (dm^3)

$$= 2 \times \frac{500}{1000}$$
$= 1$

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

mass of 1 mole of KOH
$= (1 \times 39) + (1 \times 16) + (1 \times 1)$
$= 56\,g$

Therefore:

mass of KOH in 1 mole
= number of moles × mass of 1 mole
$= 1 \times 56$
$= 56\,g$

Questions

Use the values of A_r 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 $24\,dm^3$.

1 Calculate the number of moles in:
   a 2 g of neon atoms
   b 4 g of magnesium atoms
   c 24 g of carbon atoms.
2 Calculate the mass of:
   a 0.1 mole of oxygen molecules
   b 5 moles of sulfur atoms
   c 0.25 mole of sodium atoms.
3 Calculate the number of moles in:
   a 9.8 g of sulfuric acid (H_2SO_4)
   b 40 g of sodium hydroxide (NaOH)
   c 720 g of iron(II) oxide (FeO).
4 Calculate the mass of:
   a 2 moles of zinc oxide (ZnO)
   b 0.25 mole of hydrogen sulfide (H_2S)
   c 0.35 mole of copper(II) sulfate ($CuSO_4$).
5 Calculate the number of moles at rtp in:
   a $2\,dm^3$ of carbon dioxide (CO_2)
   b $240\,dm^3$ of sulfur dioxide (SO_2)
   c $20\,cm^3$ of carbon monoxide (CO).
6 Calculate the volume of:
   a 0.3 mole of hydrogen chloride (HCl)
   b 4.4 g of carbon dioxide
   c 34 g of ammonia (NH_3).
7 Calculate the concentration of solutions containing:
   a 0.2 mole of sodium hydroxide dissolved in water and made up to $100\,cm^3$
   b 9.8 g of sulfuric acid dissolved in water and made up to $500\,cm^3$.
8 Calculate the mass of:
   a copper(II) sulfate ($CuSO_4$) which needs to be used to prepare $500\,cm^3$ of a $0.1\,mol\,dm^{-3}$ solution
   b potassium nitrate (KNO_3) which needs to be used to prepare $200\,cm^3$ of a $2\,mol\,dm^{-3}$ 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 $PbBr_2$. 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).

magnesium + oxygen → magnesium oxide

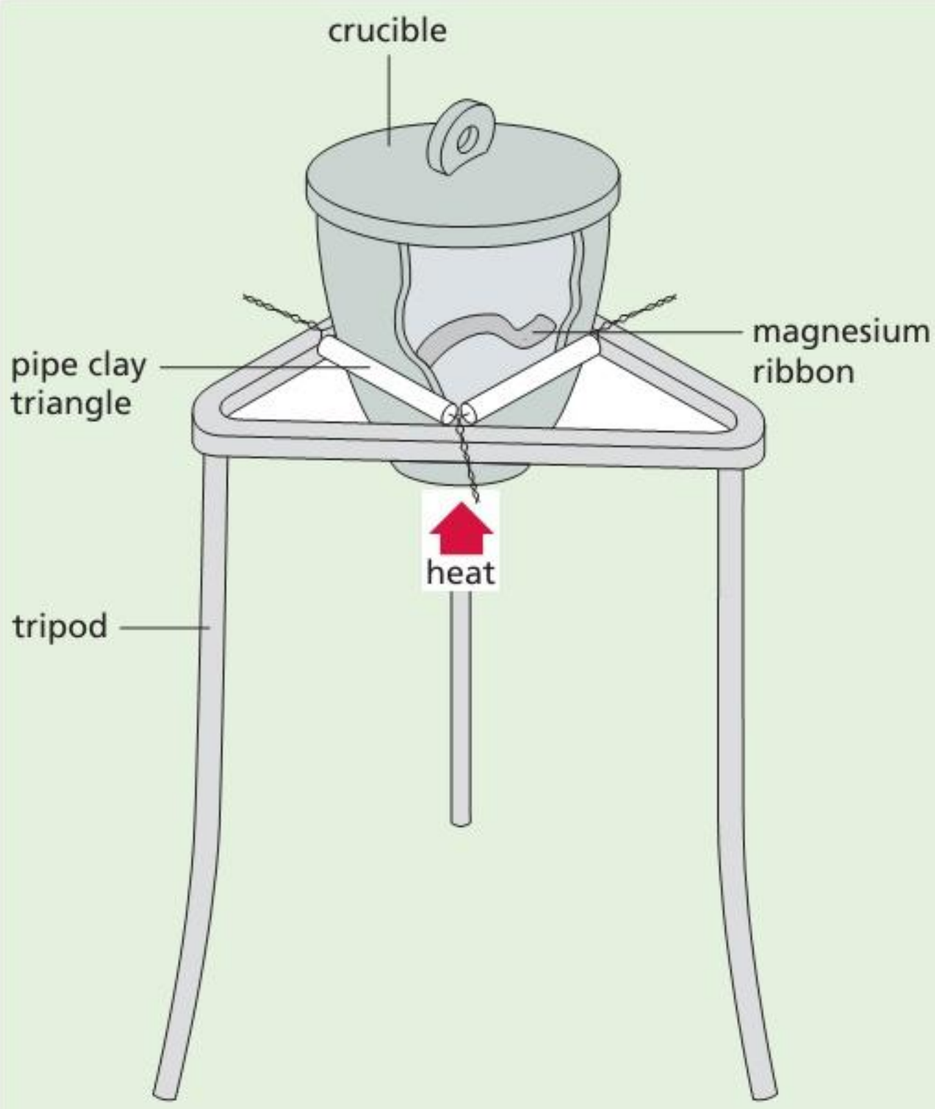


Figure 4.5 Apparatus used to determine magnesium oxide's formula.

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)

	Mg	**O**
Masses reacting (g)	0.24	0.16
Number of moles	$\frac{0.24}{24}$ = 0.01	$\frac{0.16}{16}$ = 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)

	C	**H**
Masses (g)	0.12	0.02
Number of moles	$\frac{0.12}{12}$ = 0.01	$\frac{0.02}{1}$ = 0.02
Ratio of moles	1	2
Empirical formula	CH_2	

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_2H_4, C_3H_6, C_4H_8 and C_5H_{10}. All of these formulae show the same ratio of carbon atoms to hydrogen atoms, CH_2, 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:

M_r of the empirical formula unit
$= (1 \times 12) + (2 \times 1)$
$= 14$

Number of empirical formula units present

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

$$= \frac{56}{14}$$

$= 4$

Therefore, the actual formula of the unknown organic compound is $4 \times CH_2 = C_4H_8$.

This substance is called butene. C_4H_8 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)

	C	H
% by mass	92.3	7.7
Masses in 100 g	92.3 g	7.7 g
Number of moles	$\frac{92.3}{12}$ = 7.7	$\frac{7.7}{1}$ = 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 × CH = C_6H_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.

1. Determine the empirical formula of an oxide of calcium formed when 0.4 g of calcium reacts with 0.16 g of oxygen.
2. 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.

magnesium + oxygen → magnesium oxide

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

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.

$2Mg(s)$	+	$O_2(g)$	→	$2MgO(s)$
2 moles		1 mole		2 moles
2 × 24 = 48 g		1 × (16 × 2) = 32 g		2 × (24 + 16) = 80 g

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, $CaCO_3$).

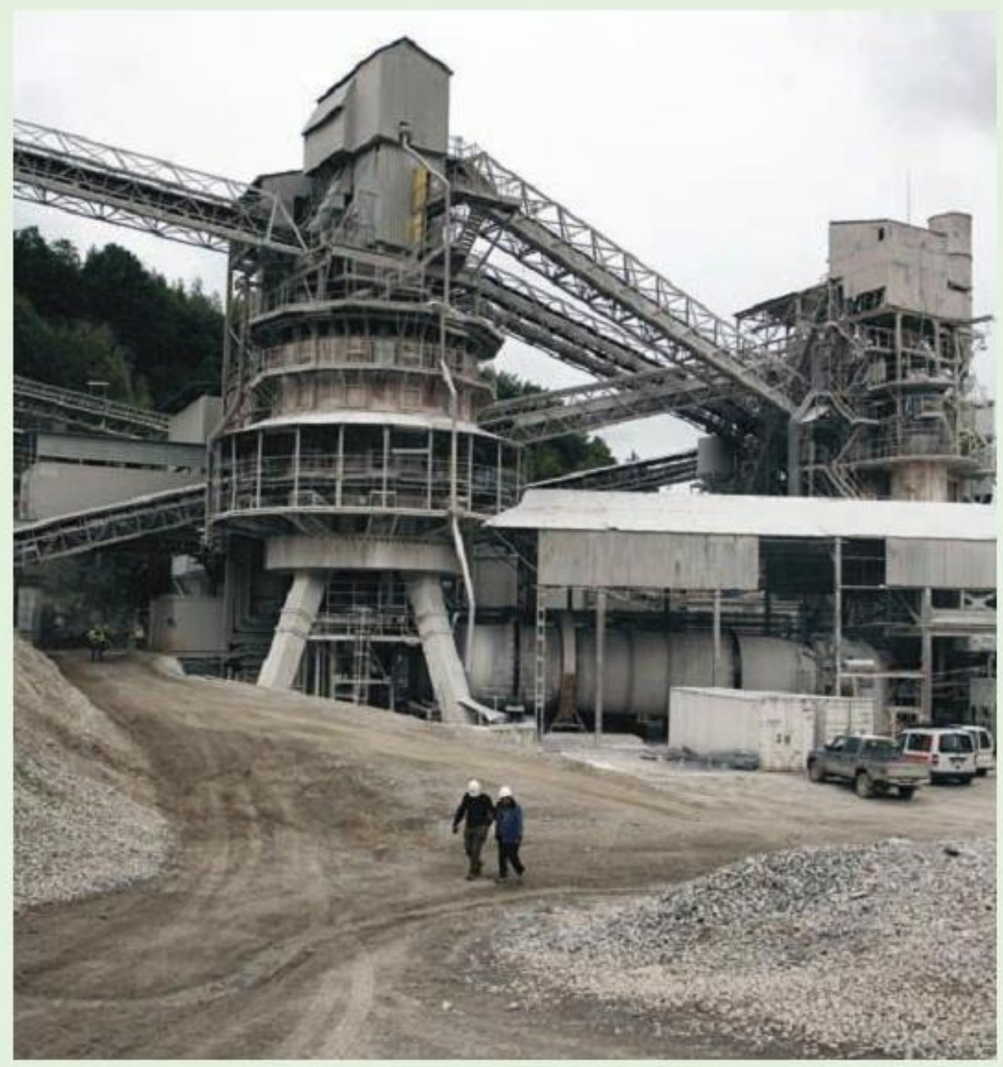


Figure 4.6 A rotary kiln for burning (calcining) limestone into lime, located in Moha, Belgium.

The equation for the process is:

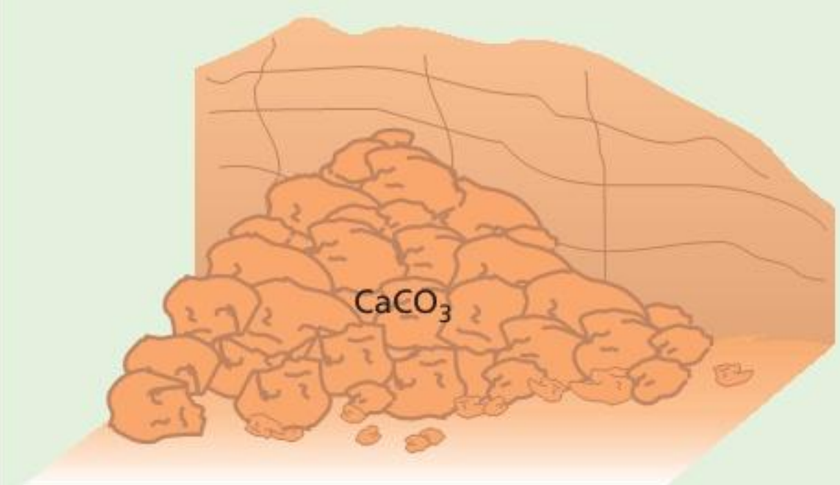
$CaCO_3(s)$	$\rightarrow$	$CaO(s)$	+	$CO_2(g)$
1 mole		1 mole		1 mole
$(40 + 12 + (3 \times 16))$		$40 + 16$		$12 + (2 \times 16)$
$= 100\,g$		$= 56\,g$		$= 44\,g$

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

1 tonne (t) = 1000 kg
1 kg = 1000 g

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

	$CaCO_3(s)$	$\rightarrow$	$CaO(s)$	+	$CO_2(g)$
Hence	100 t		56 t		44 t
	10 t		5.6 t		4.4 t



10 tonnes of limestone

? 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** $O_2(g)$ used and **b** water, $H_2O(g)$, produced if 960 dm^3 of hydrogen gas, $H_2(g)$, were burned in oxygen. (A_r: H = 1; O = 16) Assume 1 mole of any gas occupies a volume of 24 dm^3.

$2H_2(g)$	+	$O_2(g)$	$\rightarrow$	$2H_2O(g)$
2 moles		1 mole		2 moles
2×24		1×24		2×24
$= 48\,dm^3$		$= 24\,dm^3$		$= 48\,dm^3$

Therefore:

(× 2)	96 dm^3	48 dm^3	96 dm^3
(× 10)	960 dm^3	480 dm^3	960 dm^3

When 960 dm^3 of hydrogen are burned in oxygen:

a 480 dm^3 of oxygen are required and
b 960 dm^3 of $H_2O(g)$ 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 mol dm^{-3} solution of H_2SO_4 required to react completely with 6 g of magnesium. (A_r: Mg = 24)

number of moles of magnesium

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

$$= \frac{6}{24}$$

$$= 0.25$$

$Mg(s)$	+	$H_2SO_4(aq)$	$\rightarrow$	$MgSO_4(aq)$	+	$H_2(g)$
1 mole		1 mole		1 mole		1 mole
0.25 mol		0.25 mol		0.25 mol		0.25 mol

So 0.25 mol of $H_2SO_4(aq)$ is required. Using:

volume of $H_2SO_4(aq)$ (dm^3)

$$= \frac{\text{moles of } H_2SO_4}{\text{concentration of } H_2SO_4 \text{ (mol dm}^{-3})}$$

$$= \frac{0.25}{1}$$

$= 0.25\ dm^3$ or $250\ cm^3$

Example 4 – using a solution

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

number of moles of HCl used
= concentration (mol dm^{-3}) × volume (dm^3)
= 0.2 × 0.04
= 0.008

$HCl(aq)$	+	$NaOH(aq)$	→	$NaCl(aq)$	+	$H_2O(l)$
1 mole		1 mole		1 mole		1 mole
0.008 mol		0.008 mol		0.008 mol		0.008 mol

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

concentration of NaOH (mol dm^{-3})

$$= \frac{\text{number of moles of NaOH}}{\text{volume of NaOH (dm}^3)}$$

$$\left(\text{volume of NaOH in dm}^3 = \frac{20}{1000} = 0.02\right)$$

$$= \frac{0.008}{0.02}$$

$= 0.4\ mol\ dm^{-3}$

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.

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

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 CO_2, then the amount of carbon dioxide expected is 44 grams, 1 mole of 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.

$$2C(s) + O_2(g) \rightarrow 2CO(g)$$

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.

$$\text{percentage yield of carbon dioxide} = \frac{28}{44} \times 100$$

$$= 63.6\%$$

In general:

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

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.

percentage purity

$$= \frac{\text{mass of the pure product}}{\text{mass of the impure product obtained}} \times 100\%$$

Sodium hydrogencarbonate, $NaHCO_3$, 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 dm^3 of carbon dioxide gas was collected at room temperature and pressure (rtp).

The equation for the reaction is:

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$$

2 moles 1 mole

Step 1: Calculate the RFM of sodium hydrogencarbonate (A_r: Na = 23; C = 12; O = 16; H = 1)
RFM of $NaHCO_3$ = 84

Step 2: 2 moles of $NaHCO_3$ produces 1 mole of CO_2.
168 g of $NaHCO_3$ would give 44 g of CO_2, which would have a volume of 24 dm^3 at rtp.
84 g of $NaHCO_3$ should give 12 dm^3 of CO_2 at rtp.
Hence the mass of $NaHCO_3$ in the sample was $84 \times \frac{11.5}{12} = 80.5$ 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 dm^3 of water ($H_2O(g)$).
5. Calculate the volume of 2 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 cm^3 of 0.2 mol dm^{-3} solution of hydrochloric acid just neutralised 15 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**

  mass of compound (in grams) = number of moles × 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**

  mass of element (in grams) = number of moles × 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 dm^3 (litres) at room temperature and pressure (rtp).

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

- **Calculating moles of solutions**

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

  number of moles = concentration (in mol dm^{-3}) × volume of solution (in dm^3)

- **Empirical formula** A formula showing the simplest ratio of atoms present.
- **Mole** The amount of substance which contains 6×10^{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

1 Calculate the mass of:
- **a** 1 mole of:
  - **(i)** chlorine molecules
  - **(ii)** iron(III) oxide.
- **b** 0.5 mole of:
  - **(i)** magnesium nitrate
  - **(ii)** ammonia.

2 Calculate the volume occupied, at rtp, by the following gases. (One mole of any gas occupies a volume of 24 dm^3 at rtp.)
- **a** 12.5 moles of sulfur dioxide gas.
- **b** 0.15 mole of nitrogen gas.

3 Calculate the number of moles of gas present in the following:
- **a** 36 cm^3 of sulfur dioxide
- **b** 144 dm^3 of hydrogen sulfide.

4 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

5 **a** Calculate the empirical formula of an organic liquid containing 26.67% of carbon and 2.22% of hydrogen, with the rest being oxygen.
- **b** The M_r of the liquid is 90. What is its molecular formula?

6 Iron is extracted from its ore, haematite, in the blast furnace. The main extraction reaction is:

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$$

- **a** Name the reducing agent in this process.
- **b** Name the oxide of iron shown in the equation.
- **c** Explain why this is a **redox** reaction.
- **d** Calculate the mass of iron which will be produced from 640 tonnes of haematite.

7 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)
- **a** Calculate the empirical formula of vulcium chloride.
- **b** To which group of the Periodic Table (p. 136) should vulcium be assigned?
- **c** Write a word and balanced chemical equation for the reaction between vulcium and chlorine.
- **d** What other information in the description supports the assignment of group you have given to vulcium?
- **e** What type of bonding is present in vulcium chloride?
- **f** Write a word and balanced chemical equation for the reaction between vulcium and water.
- **g** Write the formulae for:
  - **(i)** vulcium sulfate
  - **(ii)** vulcium carbonate
  - **(iii)** vulcium hydroxide.

Look at the Periodic Table (p. 294) to find out the real name of vulcium.

8 0.048 g of magnesium was reacted with excess dilute hydrochloric acid at room temperature and pressure. The hydrogen gas given off was collected.
- **a** Write a word and balanced symbol equation for the reaction taking place.
- **b** Draw a diagram of an apparatus which could be used to carry out this experiment and collect the hydrogen gas.
- **c** How many moles of magnesium were used?
- **d** 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.
- **e** Calculate the volume of a solution containing 0.1 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).

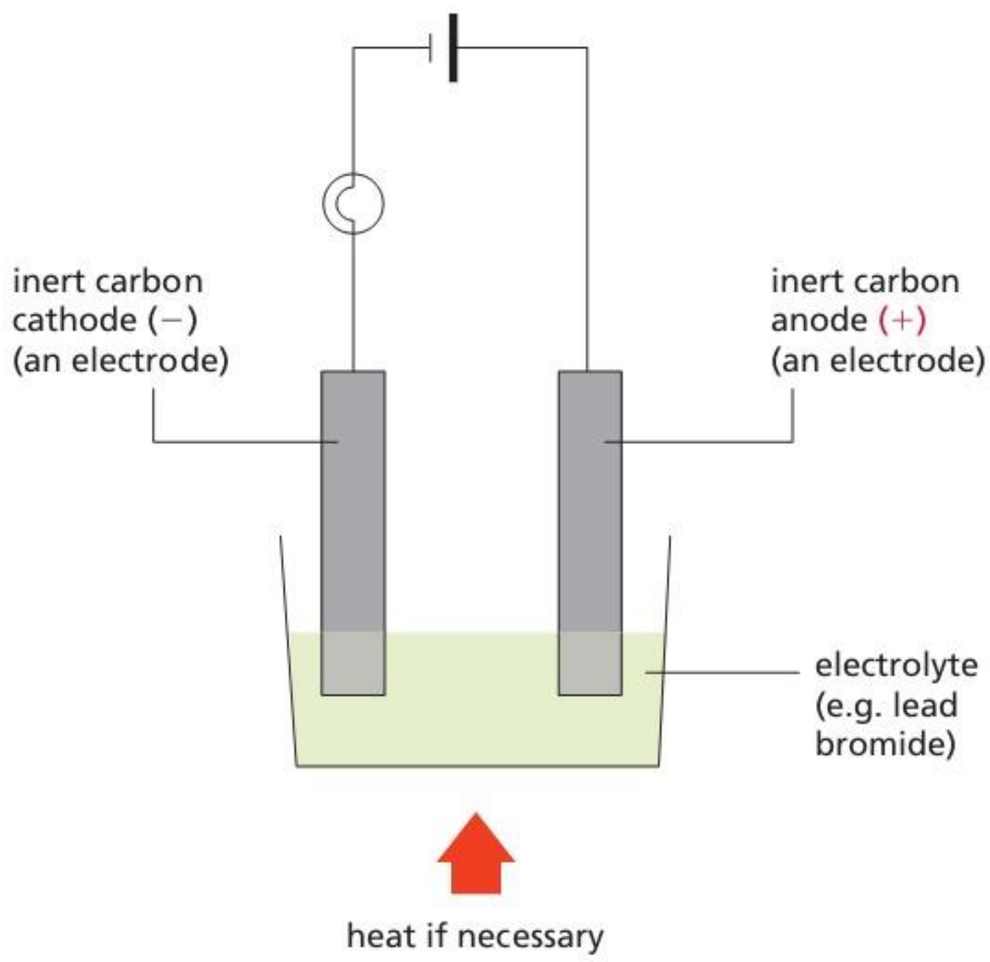


Figure 5.2 The important terms used in electrolysis.

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 ($PbBr_2$) 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**.

$$\text{molten lead(II) bromide} \rightarrow \text{bromine} + \text{lead}$$
$$PbBr_2(l) \rightarrow Br_2(g) + Pb(l)$$

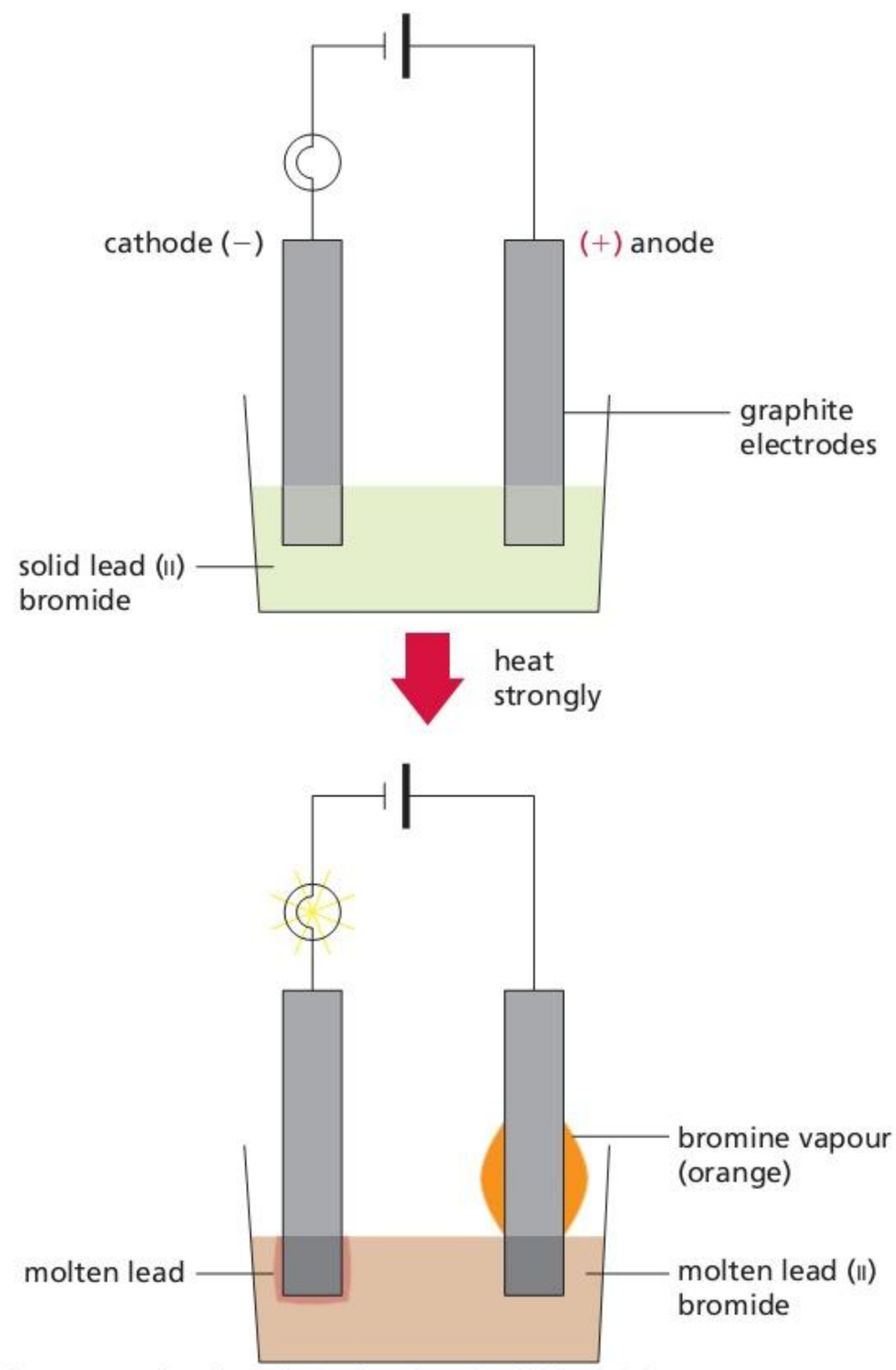


Figure 5.3 The electrolysis of molten lead(II) bromide.

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:

$$\text{lead ion} + \text{electrons} \rightarrow \text{lead atom}$$
$$Pb^{2+}(l) + 2e^- \rightarrow Pb(l)$$

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:

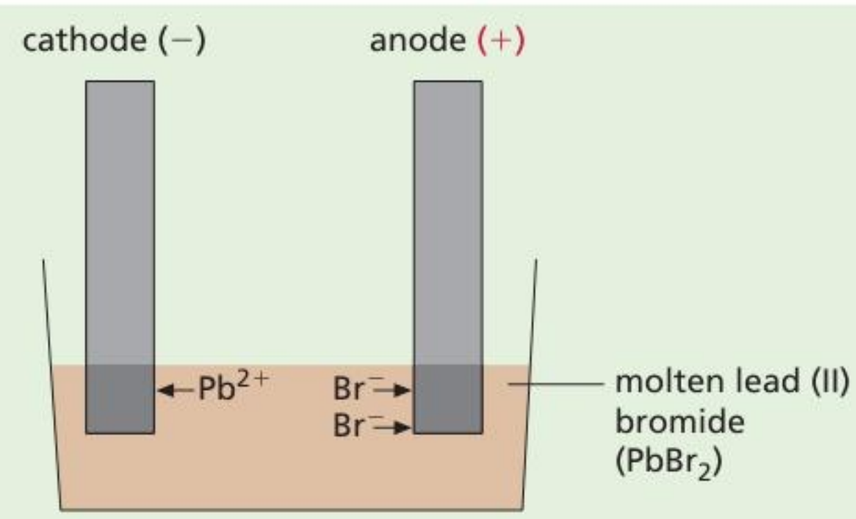


Figure 5.4 The lead ions (Pb^{2+}) are attracted to the cathode and the bromide ions (Br^-) are attracted to the anode.

$$\text{bromide ion} \rightarrow \text{bromine atom} + \text{electron}$$
$$Br^-(l) \rightarrow Br + e^-$$

Two bromine atoms then combine to form a bromine molecule:

$$\text{bromine atoms} \rightarrow \text{bromine molecule}$$
$$2Br \rightarrow Br_2(g)$$

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 (Na_3AlF_6). Cryolite, a mineral found naturally in Greenland, is used to reduce the working temperature of the Hall–Héroult cell from 2017 °C (the melting point of pure aluminium oxide) to between 800 and 1000 °C. Therefore, the cryolite provides a considerable

Figure 5.5 Bauxite mining.

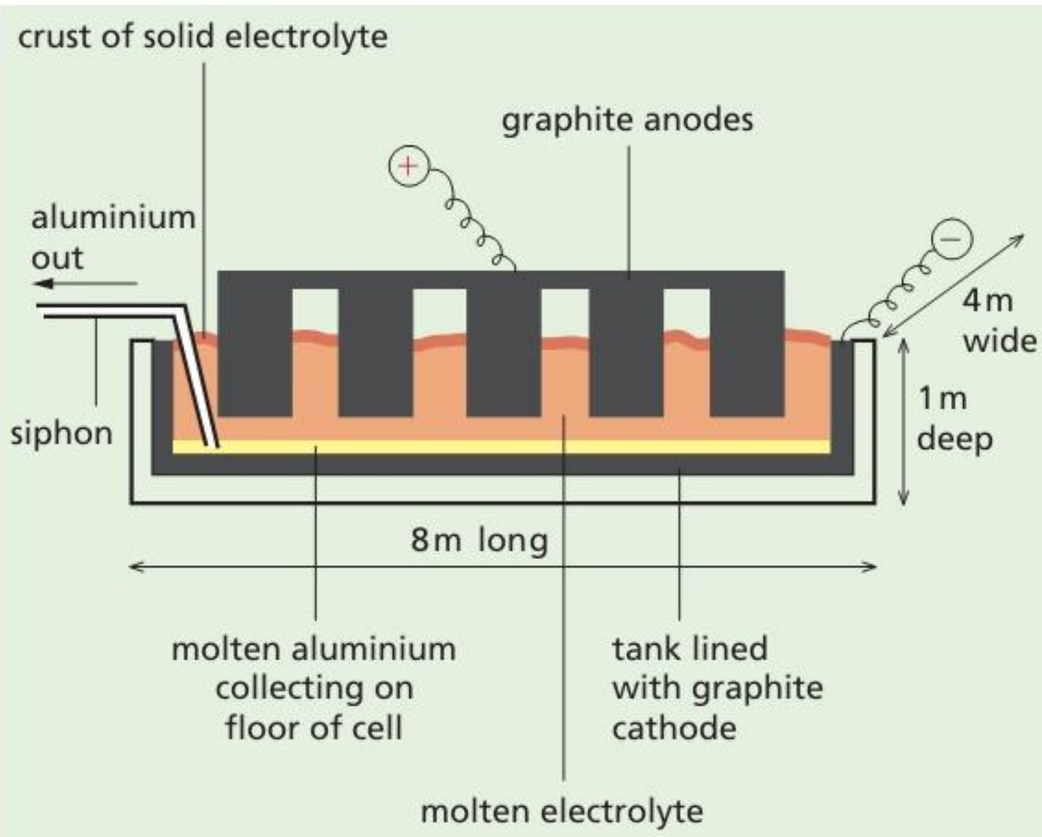


Figure 5.6 The Hall–Héroult cell is used in industry to extract aluminium by electrolysis.

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).

oxide ions → oxygen molecules + electrons

$$2O^{2-}(l) \rightarrow O_2(g) + 4e^-$$

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

aluminium ions + electrons → aluminium metal

$$Al^{3+}(l) + 3e^- \rightarrow Al(l)$$

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:

aluminium oxide $\xrightarrow{\text{electrolysis}}$ aluminium + oxygen

$$2Al_2O_3(l) \longrightarrow 4Al(l) + 3O_2(g)$$

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.

carbon + oxygen → carbon dioxide

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

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

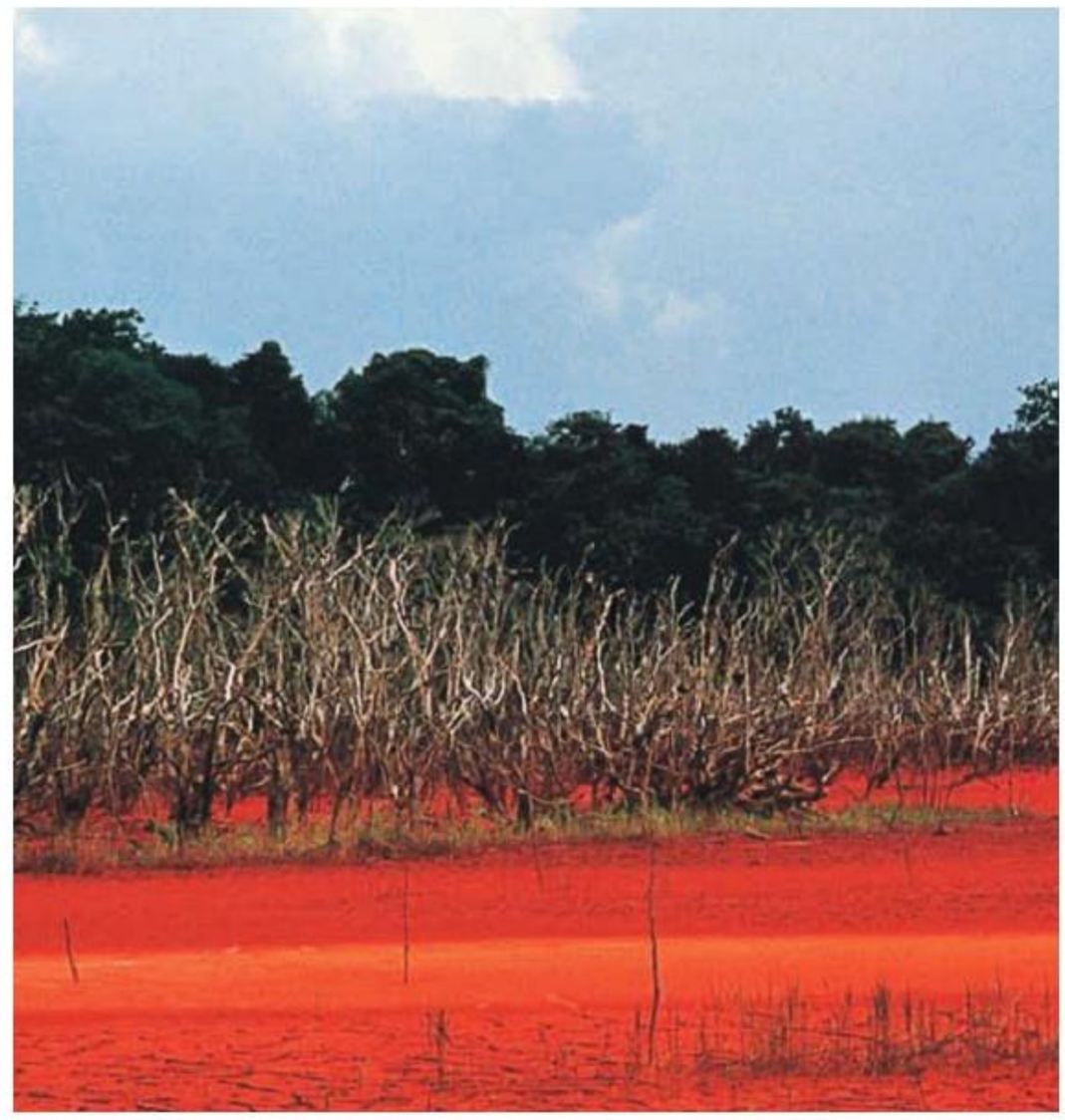


Figure 5.9 Bauxite pollution of the Amazon in Brazil.

- 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 (Al_2O_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.

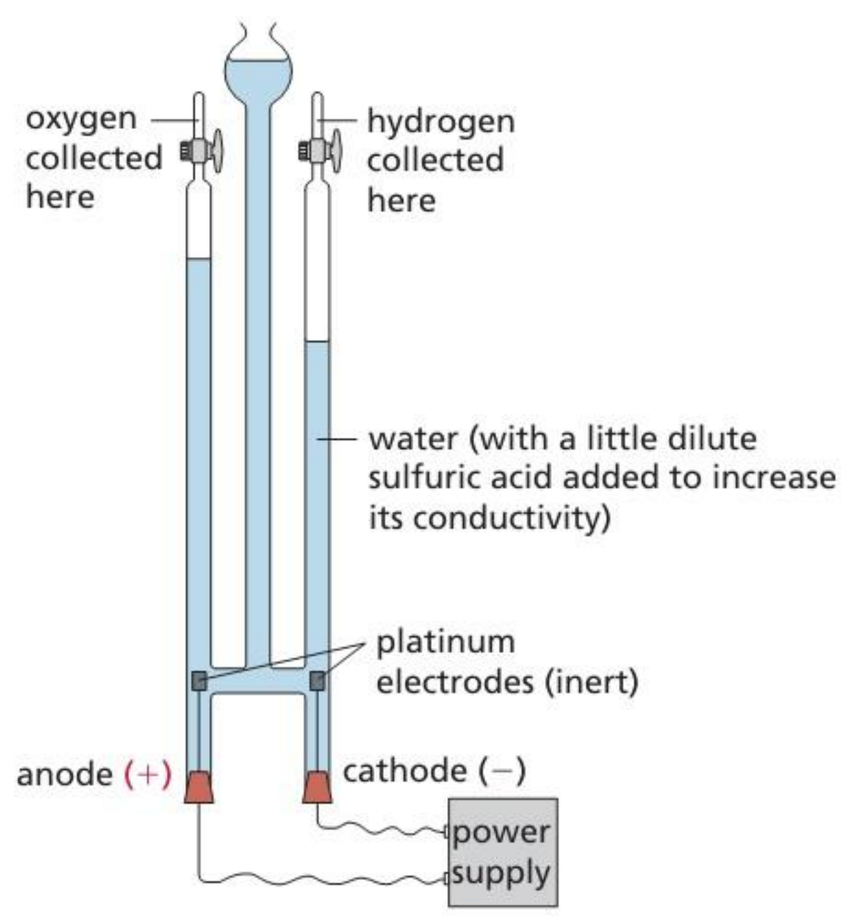


Figure 5.11 A Hofmann voltameter used to electrolyse water.

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.

hydrogen ions + electrons → hydrogen molecules

$$4H^+(aq) + 4e^- \rightarrow 2H_2(g)$$

If during this process the water molecules lose $H^+(aq)$, then the remaining portion must be hydroxide ions, $OH^-(aq)$. 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.

hydroxide ions → water molecules + oxygen molecules + electrons

$$4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$$

This experiment was first carried out by Sir Humphry Davy. It confirmed that the formula for water was H_2O.

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 electrolysing 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:

from the water:	$H^+(aq)$	$OH^-(aq)$
from the sodium chloride:	$Na^+(aq)$	$Cl^-(aq)$

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

chloride ions $\xrightarrow{\text{oxidation}}$ chlorine molecules + electrons

$$2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-$$

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

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

hydrogen ions + electrons $\xrightarrow{\text{reduction}}$ hydrogen molecules

$$2H^+(aq) + 2e^- \longrightarrow H_2(g)$$

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

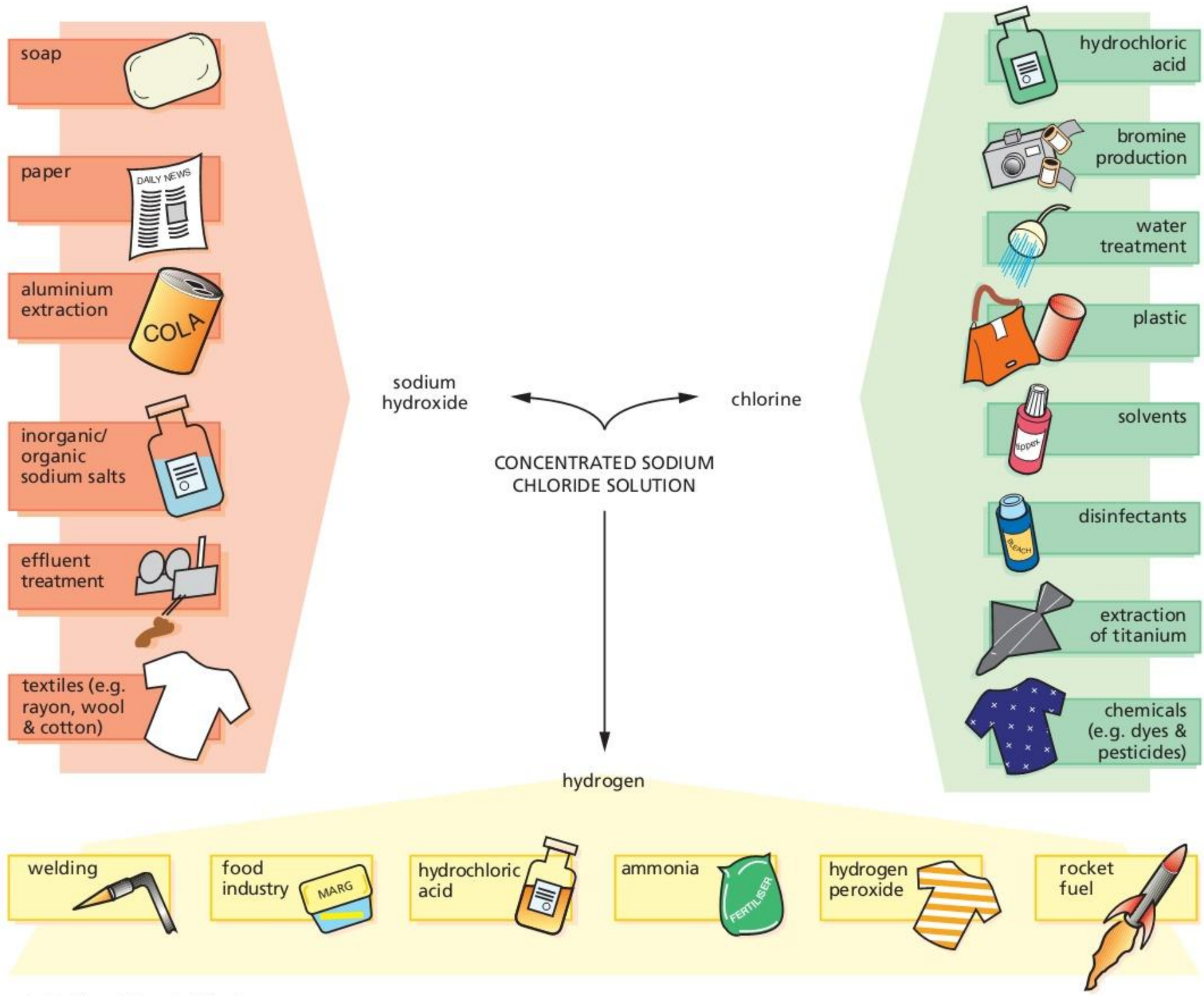
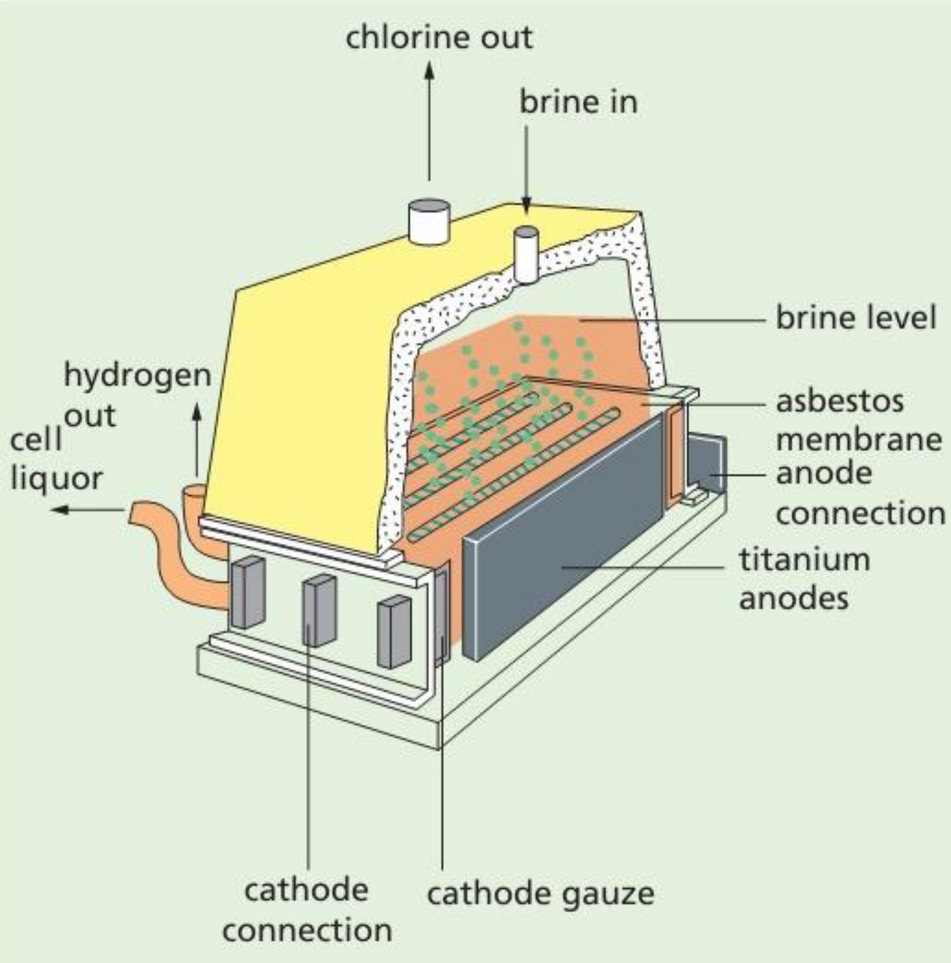
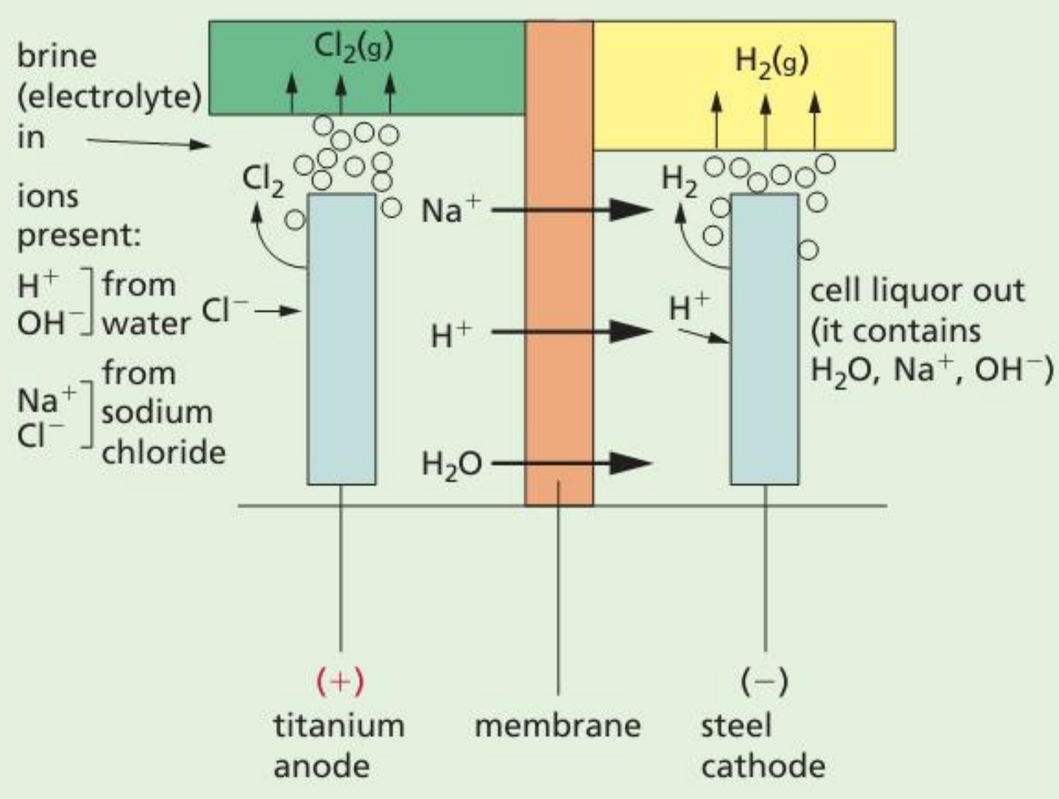


Figure 5.12 The chlor-alkali industry.



a A section through the membrane cell.



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

Figure 5.13

Tests for aqueous cations

Aqueous sodium hydroxide can be used to identify salts of Al^{3+}, Ca^{2+}, Cr^{3+}, Cu^{2+}, Fe^{2+}, Fe^{3+} and 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 (Al^{3+})	White precipitate, soluble in excess giving a colourless solution
Ammonium (NH_4^+)	Ammonia produced on warming
Calcium (Ca^{2+})	White precipitate, insoluble in excess
Chromium (Cr^{3+})	Green precipitate, soluble in excess
Copper (Cu^{2+})	Light blue precipitate, insoluble in excess
Iron(II) (Fe^{2+})	Green precipitate, insoluble in excess
Iron(III) (Fe^{3+})	Red-brown precipitate, insoluble in excess
Zinc (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.

Questions

1 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.

2 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.
   a The universal indicator initially turns red in the region of the anode, but as the electrolysis proceeds it loses its colour.
   b The universal indicator turns blue in the region of the cathode.

3 Why is it important to remove compounds of calcium, strontium and magnesium before brine is electrolysed?

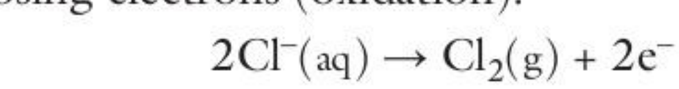
4 The uses of sodium hydroxide can be separated on a percentage basis as follows:

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 concentrated hydrochloric acid

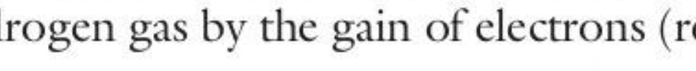
When concentrated hydrochloric acid is electrolysed, two gases are produced. Both chloride ions and hydroxide ions would be attracted to the anode. The chloride ions are at a much higher concentration than the hydroxide ions and go on to produce chlorine gas by losing electrons (oxidation).



$$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$$

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).



$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

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

Electrolysis of copper(II) sulfate solution

Copper(II) sulfate solution ($CuSO_4(aq)$) 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:	$H^+(aq)$	$OH^-(aq)$
from the copper(II) sulfate:	$Cu^{2+}(aq)$	$SO_4^{2-}(aq)$

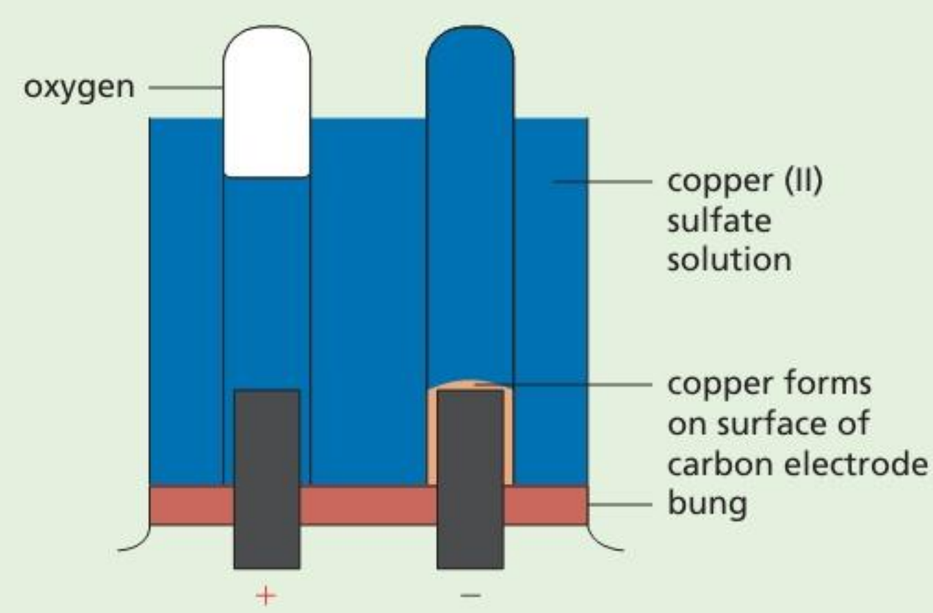


Figure 5.14 The electrolysis of copper(II) sulfate solution using inert electrodes.

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

copper ions + electrons → copper atoms

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

$OH^-(aq)$ and $SO_4^{2-}(aq)$ ions are both attracted to the anode. The OH^- ions release electrons more easily than the SO_4^{2-} ions, so oxygen gas and water are produced at the anode (Figure 5.15).

hydroxide ions → oxygen + water + electrons

$$4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$$

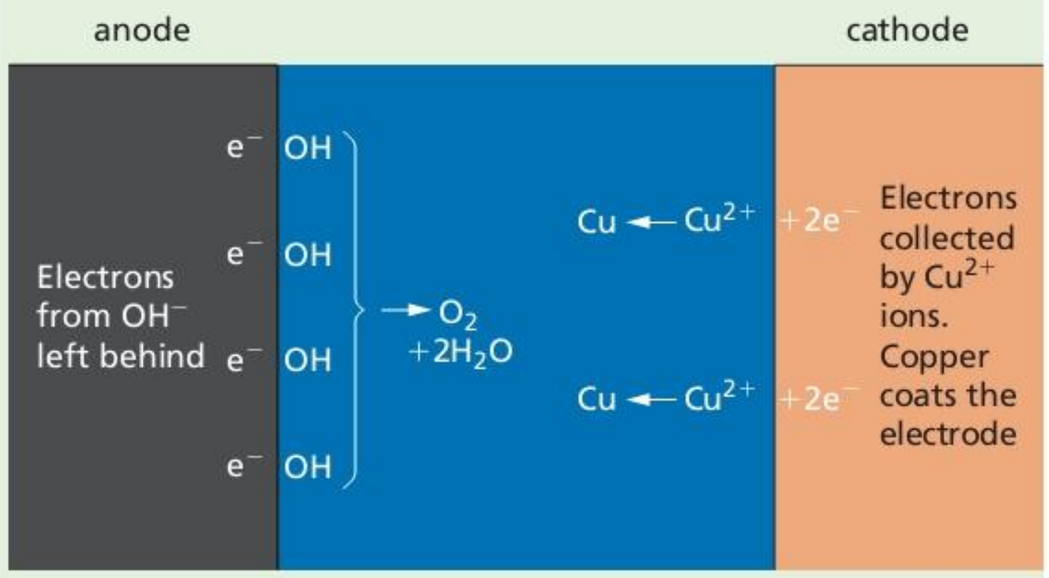


Figure 5.15 Oxygen is given off at the anode and copper is deposited at the cathode.

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.

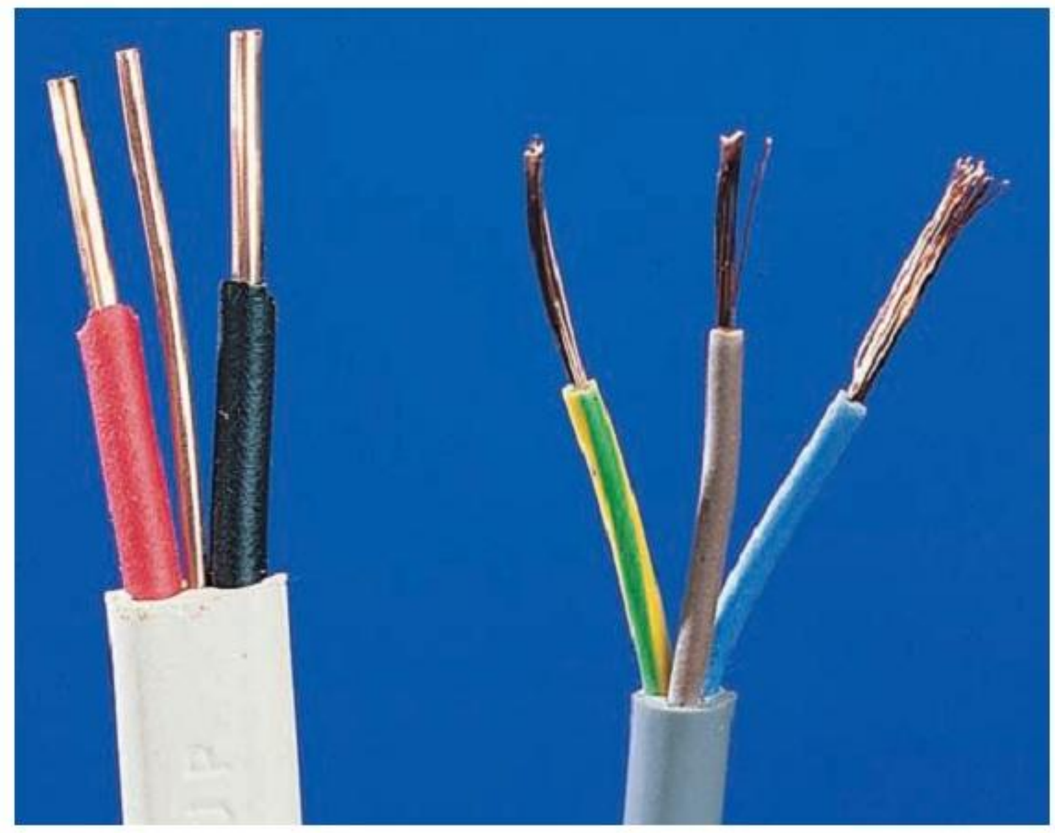


Figure 5.16a The copper used in electrical wiring has to be very pure.

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 mol dm^{-3}) acidified with a 2 mol dm^{-3} solution of sulfuric acid to help the solution conduct electricity (Figure 5.17).

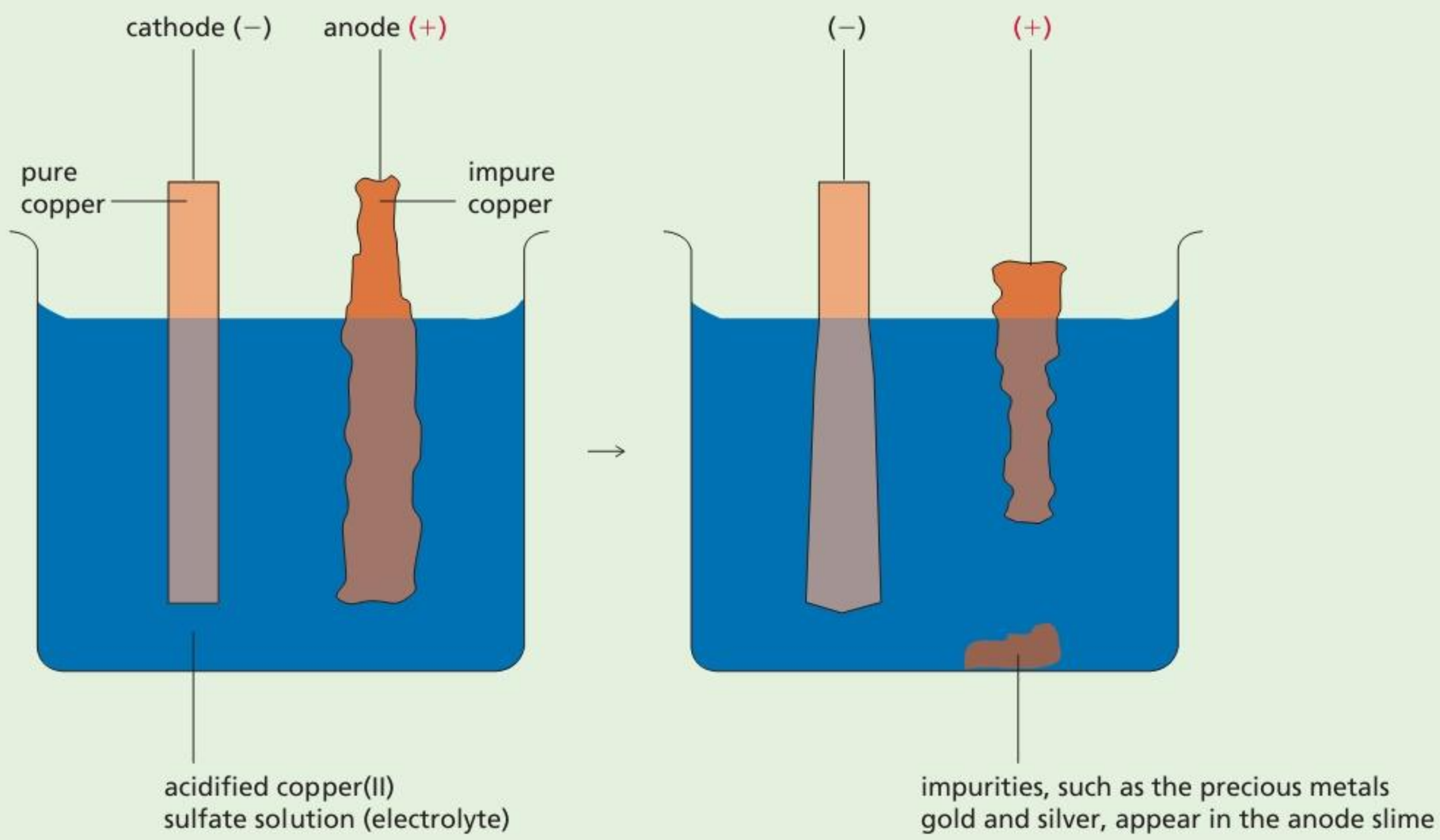


Figure 5.17 Copper purification process.

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 Am^{-2} are usually used.

The ions present in the solution are:

from the water:	$H^+(aq)$	$OH^-(aq)$
from the copper(II) sulfate:	$Cu^{2+}(aq)$	$SO_4^{2-}(aq)$

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

copper atoms → copper ions + electrons

$$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$$

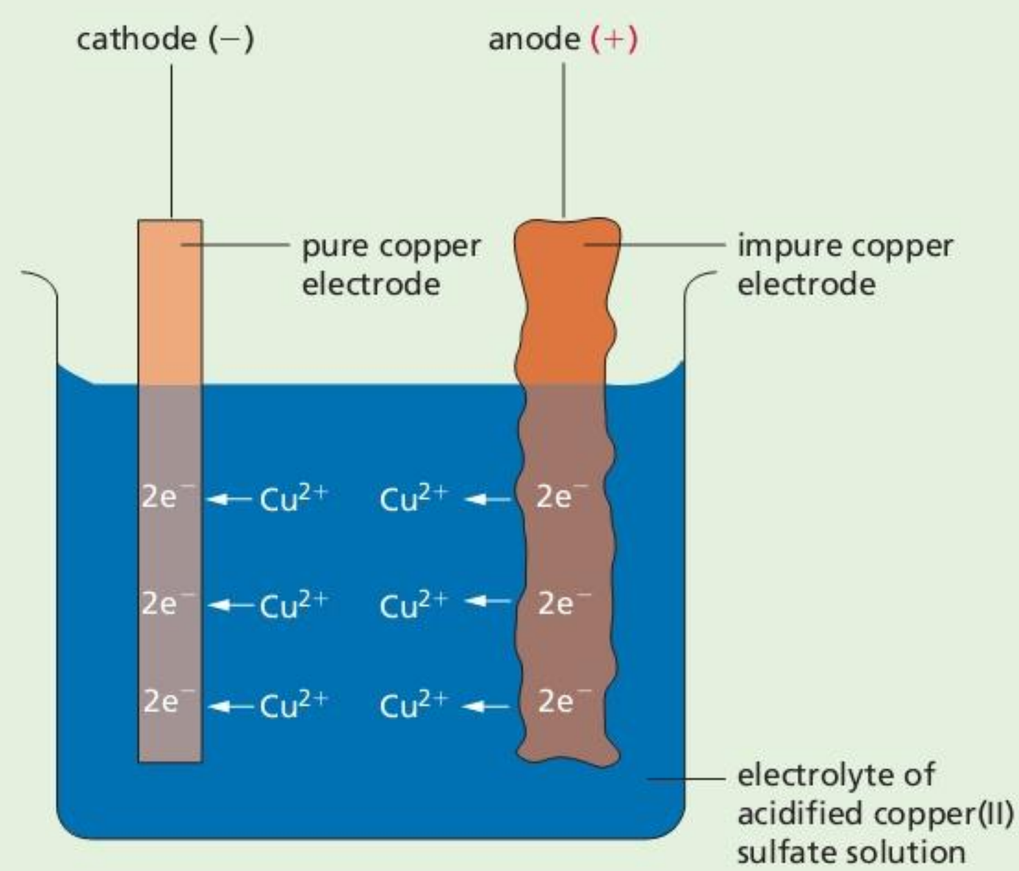


Figure 5.18 The movement of ions in the purification of copper by electrolysis.

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

copper ions + electrons → copper atoms

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

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).

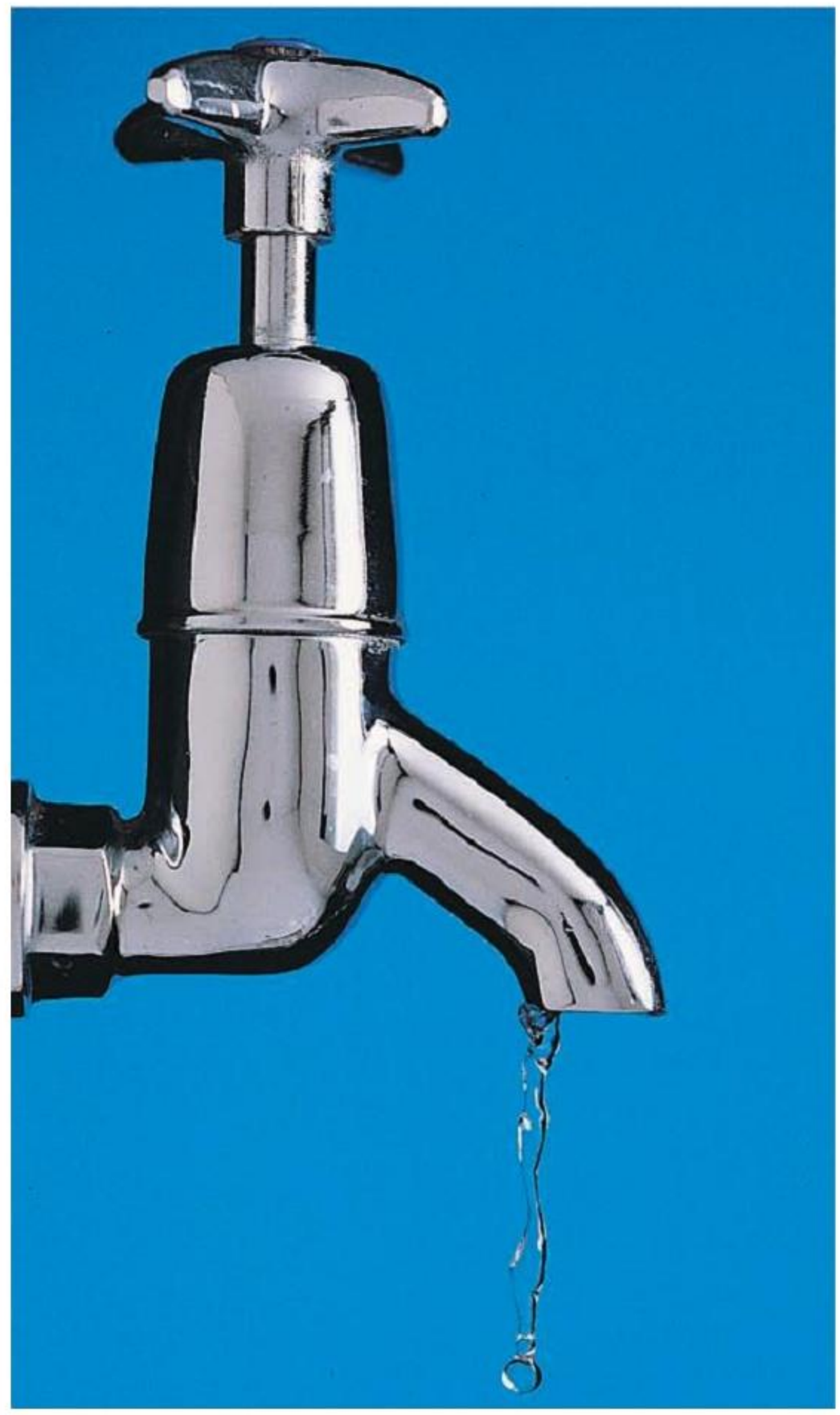


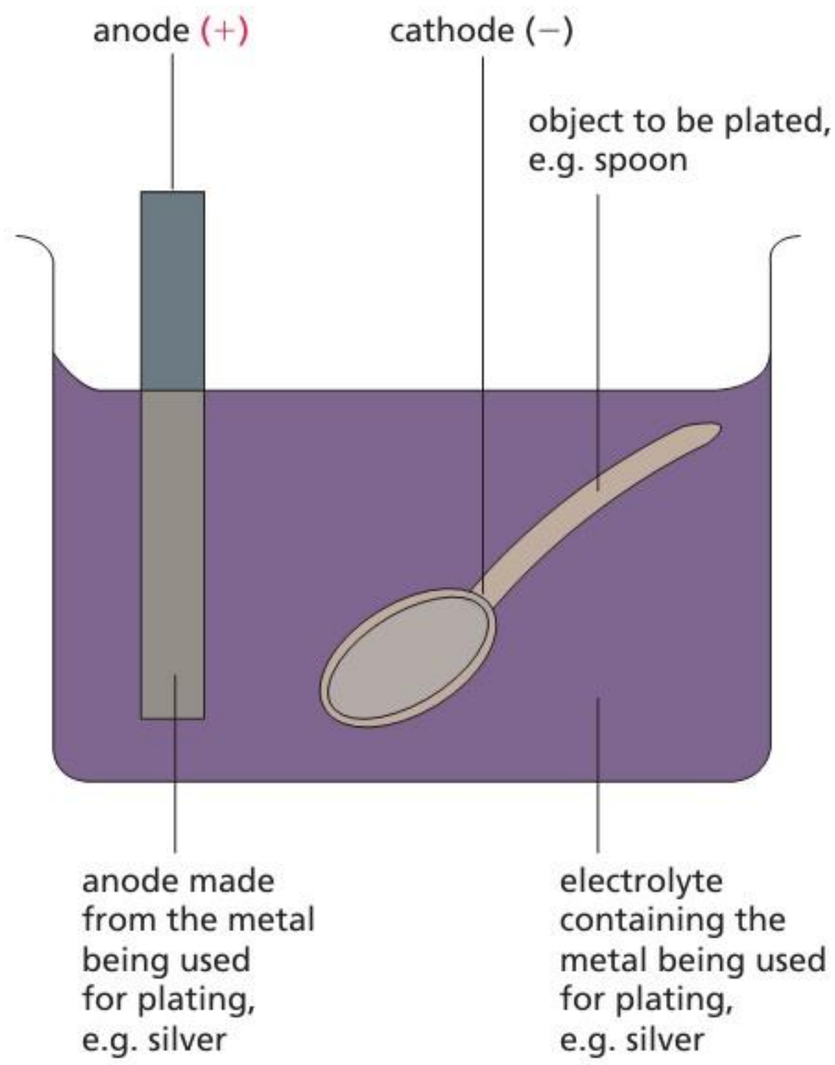
Figure 5.19 This tap has been chromium plated.

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.

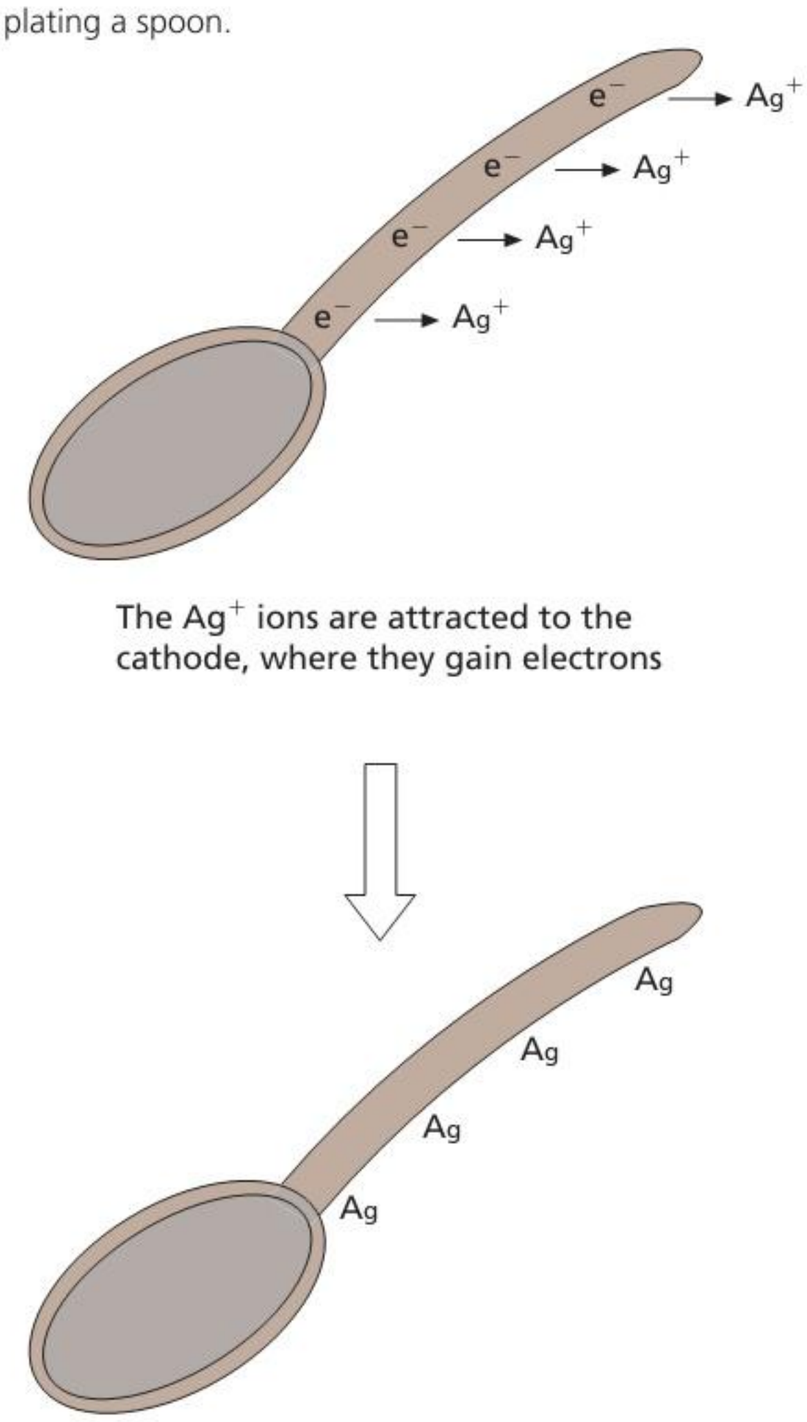
For silver plating the electrolyte is a solution of a silver salt. The article to be plated is made the cathode in the cell so that the metal ions move to it when the current is switched on. The cathode reaction in this process is:

silver ions + electrons → silver atoms

$$Ag^+(aq) + e^- \rightarrow Ag(s)$$



a Silver plating a spoon.



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

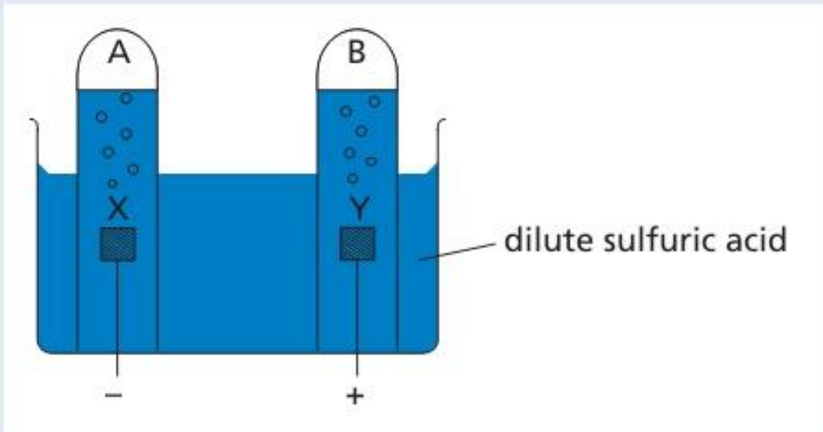
After studying Chapter 5 you should know and understand the following terms.

- **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 °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

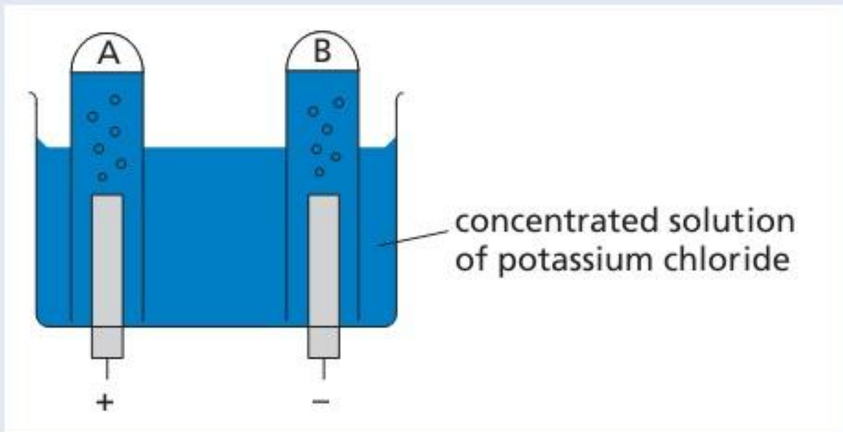
1



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

a Name the gas that collects at **A**.
b Name the gas that collects at **B**.
c If 100 cm^3 of gas collects in **A** how much would there be in **B**?
d Name the metal usually used for **X** and **Y**.
e **X** is called the _____ .
f **Y** is called the _____ .
g Write down the formulae of the three ions present in the solution.
h 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.

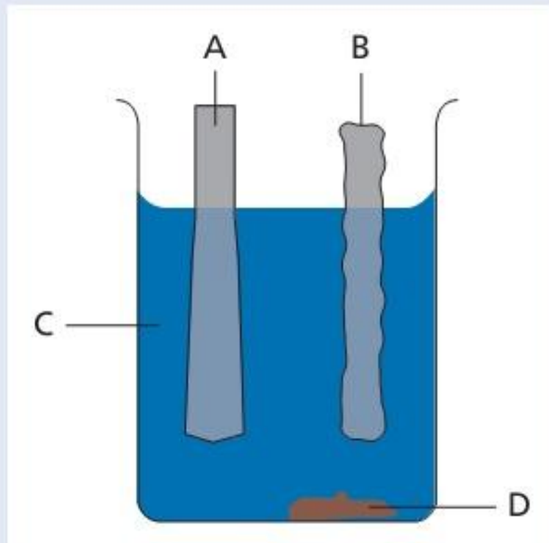


a Name a non-metal suitable for use as electrodes.
b Name the gas collected in **A** and the gas collected in **B**.
c Describe how you would test the gases collected.
d The volume of gas collected in **B** was slightly less than that collected in **A**. The teacher said the volumes should have been equal but gave a simple explanation of the 'missing' gas in **B**. What was the explanation? (Assume that the apparatus was working perfectly).
e Write down the equations which describe the production of the gases at the electrodes in **A** and **B**.
f **(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.
b Cathode.
c Electrolysis.
d Electrolyte.
e Anion.
f Cation.
g Oxidation.
h Reduction.

4 Copper is purified by electrolysis, as in the example shown below.



a Name the materials used for the electrodes **A** and **B**.
b Name the electrolyte **C** and substance **D**.
c Why is substance **D** of economic importance in respect of this process?
d Write equations for the reactions which take place at the cathode and anode during this process.
e Draw a labelled diagram to show the cell after electrolysis has taken place.
f Why has electrolyte **C** to be acidified with the dilute sulfuric acid?
g 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

a State what you understand by 'inert electrodes'.
b Explain why the lithium chloride solution becomes progressively more alkaline during electrolysis.
c Explain why solid lithium chloride is a non-conductor of electricity, whereas molten lithium chloride and lithium chloride solution are good conductors of electricity.
d During the electrolysis of molten aluminium chloride ($AlCl_3$) 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.
a Molten lead iodide.
b Sugar solution.
c Silver nitrate solution.
d Copper(II) sulfate solution.
e Molten sodium bromide.
f Ethanol solution.

7 Sodium hydroxide is made by the electrolysis of brine.
a Draw and label a simplified diagram of the cell used in this process. Make certain that you have labelled on the diagram:
(i) the electrolyte
(ii) the material of the electrodes
(iii) the material of the membrane.
b Write equations for the reactions which take place at the cathode and anode. State clearly whether a reaction is oxidation or reduction.
c Give two large-scale uses of the products of this electrolytic process.
d Comment on the following statement: 'This electrolytic process is a very expensive one'.
e 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.
a Explain what electroplating is.
b Why are certain metals electroplated?
c 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.

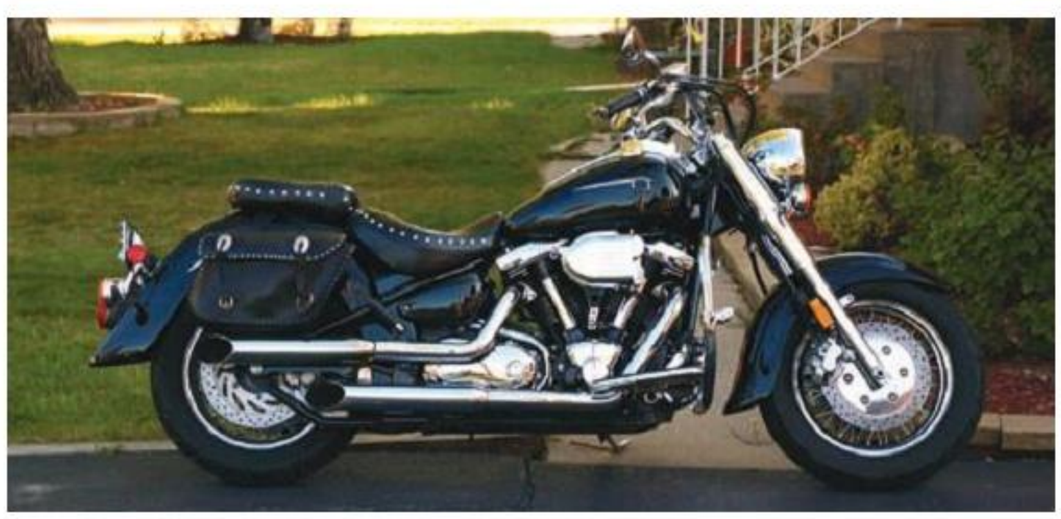


Figure 6.1 These modes of transport all use fuels that have been extracted from crude oil.

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 °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 °C and 110 °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.

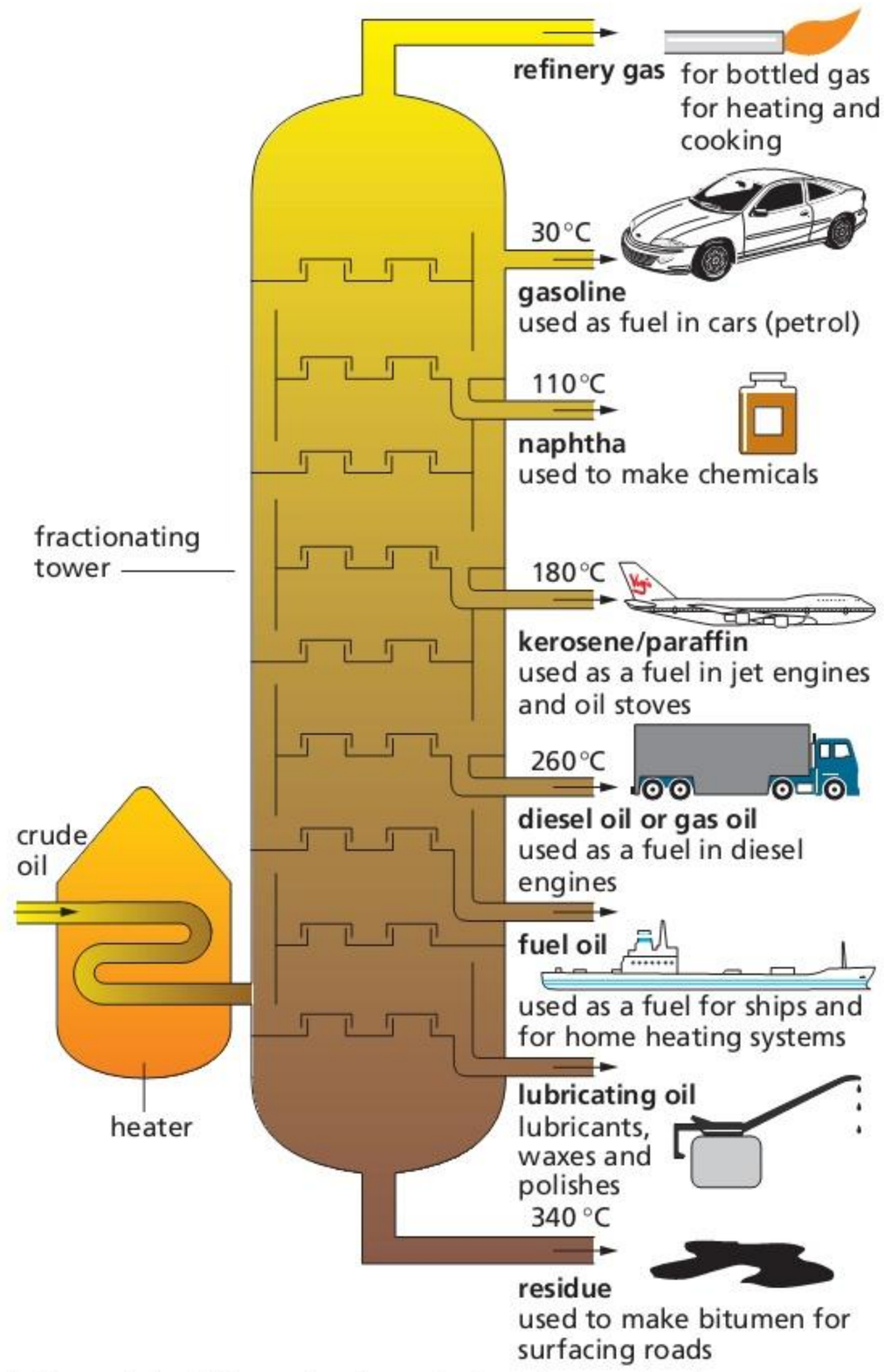
The liquids condensing at different levels are collected on **trays**. In this way the crude oil is separated into different fractions. These fractions usually contain a number of different hydrocarbons. The individual single hydrocarbons can then be obtained, again by refining the fraction by further 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 250–350 °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.



b Uses of the different fractions obtained from crude oil.

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).

Figure 6.5 Piece of coal showing a fossilised leaf.

Figure 6.6 Cutting coal today is extremely mechanised.

Figure 6.7 Oil production in the North Sea.

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).

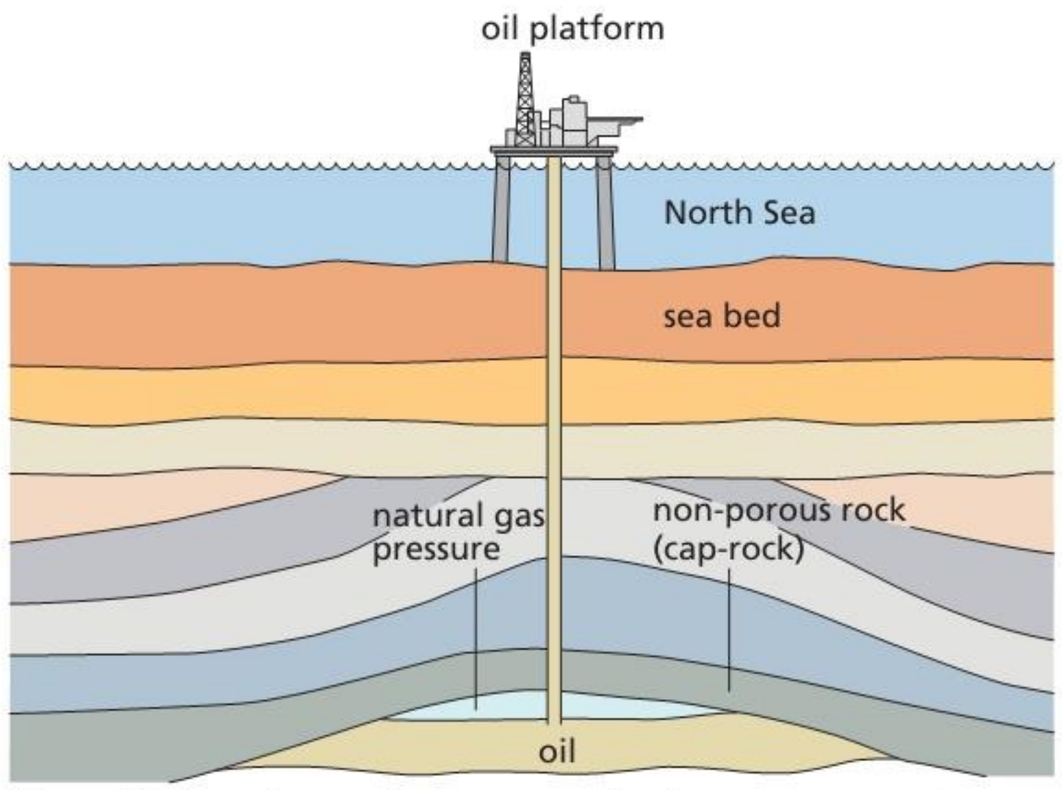


Figure 6.8 Natural gas and oil are trapped under non-porous rock.

Questions

1 Coal, oil and natural gas are all termed 'fossil fuels'. Why is the word 'fossil' used in this context?

2 a Name the process by which plants convert carbon dioxide and water into glucose.

b What conditions are necessary for this process to occur?

3 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 dioxide + water + energy

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

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.

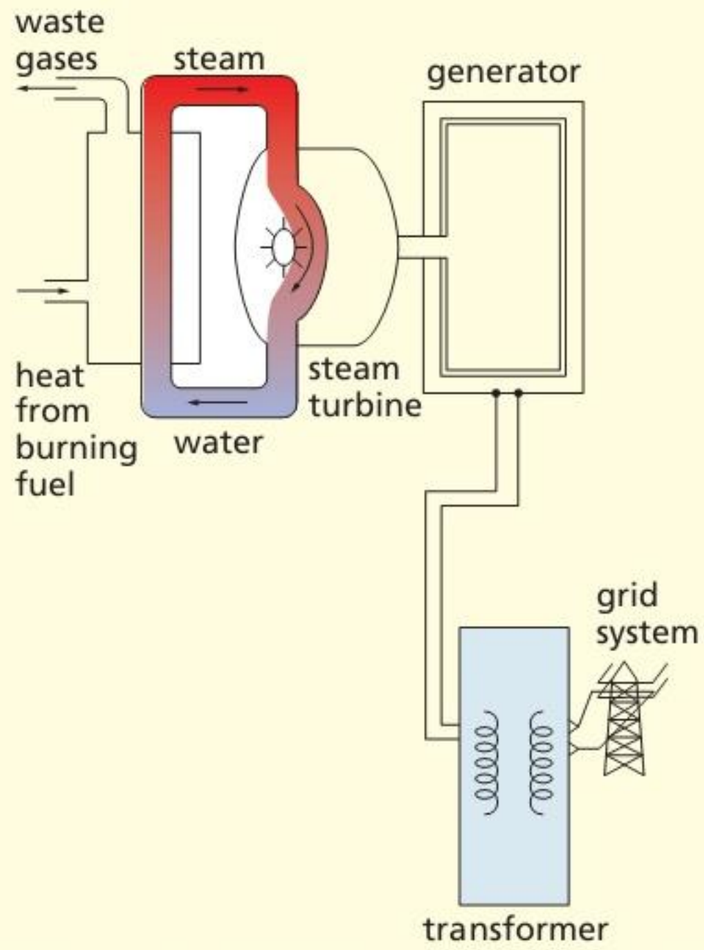


Figure 6.10 The way in which fuels are used to produce electricity.

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 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).

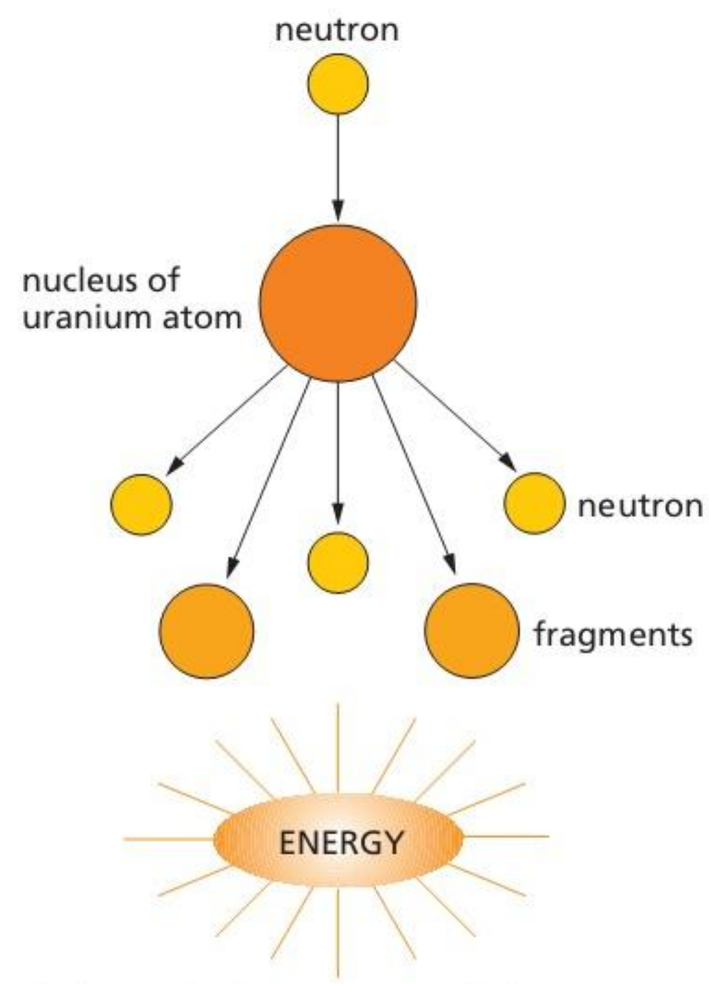


Figure 6.12 Chain reaction in uranium-235 fission.

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).

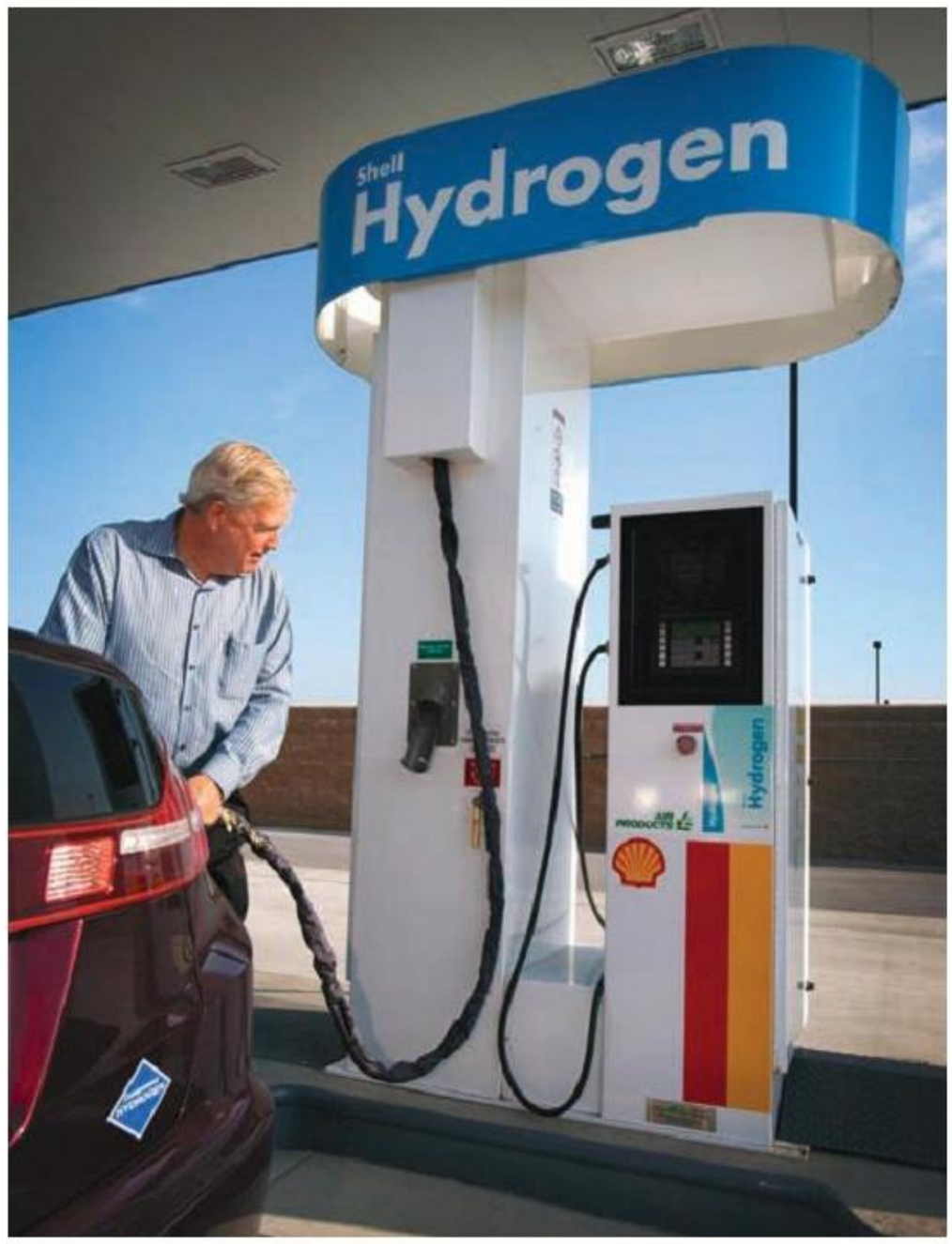


Figure 6.17 This car is powered by hydrogen gas.

Hydrogen burns cleanly to form water:

hydrogen + oxygen → water

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

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.

methane + oxygen → carbon dioxide + water + heat energy

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) + \text{heat energy}$$

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.

methane + oxygen → carbon monoxide + water + heat energy

$$2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(l) + \text{heat energy}$$

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.

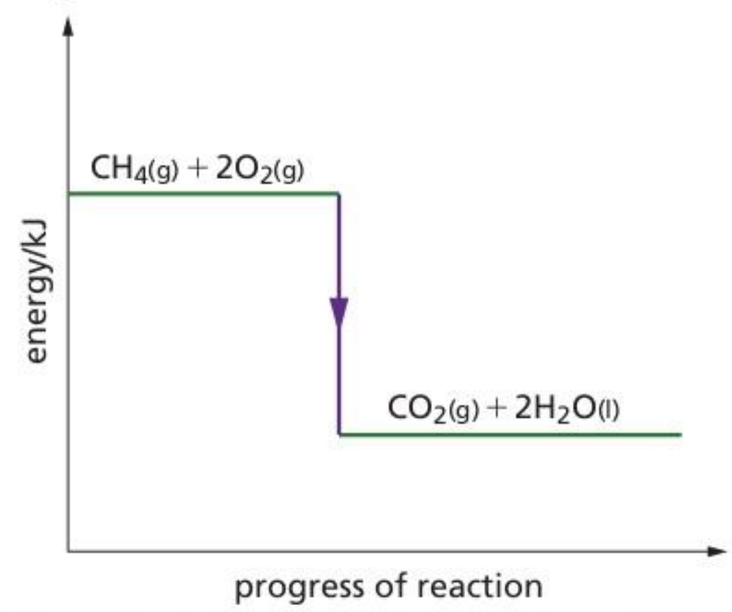


Figure 6.18 Energy level diagram for the complete combustion of methane.

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.

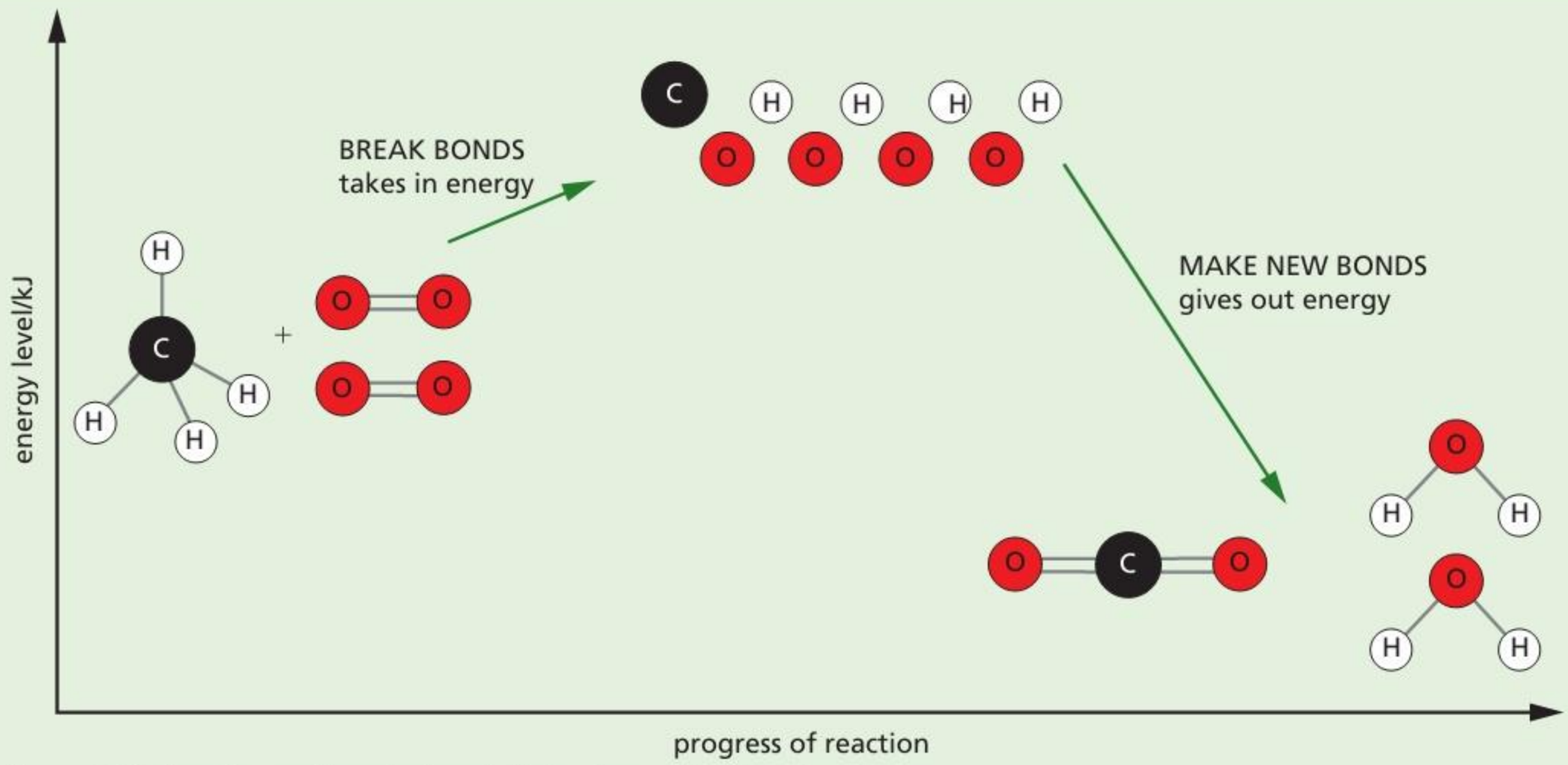


Figure 6.19 Breaking and forming bonds during the combustion of methane.

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^{-1}
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
$4 \times 435 = 1740$ kJ

Breaking 2 O=O bonds in oxygen requires
$2 \times 497 = 994$ kJ

Total = 2734 kJ of energy

Making bonds

Making 2 C=O bonds in carbon dioxide gives out
$2 \times 803 = 1606$ kJ

Making 4 O—H bonds in water gives out
$4 \times 464 = 1856$ kJ

Total = 3462 kJ of energy

Energy difference
= energy required to break bonds – energy given out when bonds are made
= 2734 – 3462
= –728 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 ΔH (pronounced 'delta H'). ΔH is called the **heat of reaction**.

For an exothermic reaction ΔH is negative and for an endothermic reaction ΔH 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.

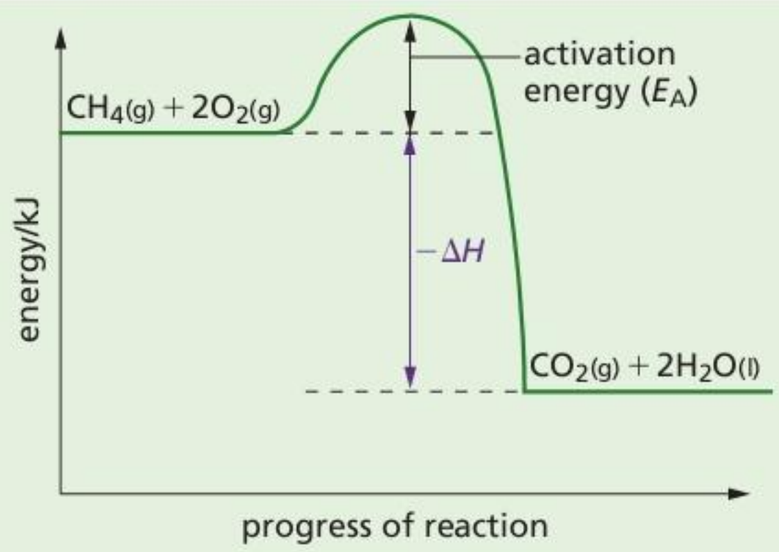


Figure 6.20 Energy level diagram for methane/oxygen.

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.

nitrogen + oxygen → nitrogen(II) oxide

$$N_2(g) + O_2(g) \rightarrow 2NO(g) + \Delta H$$

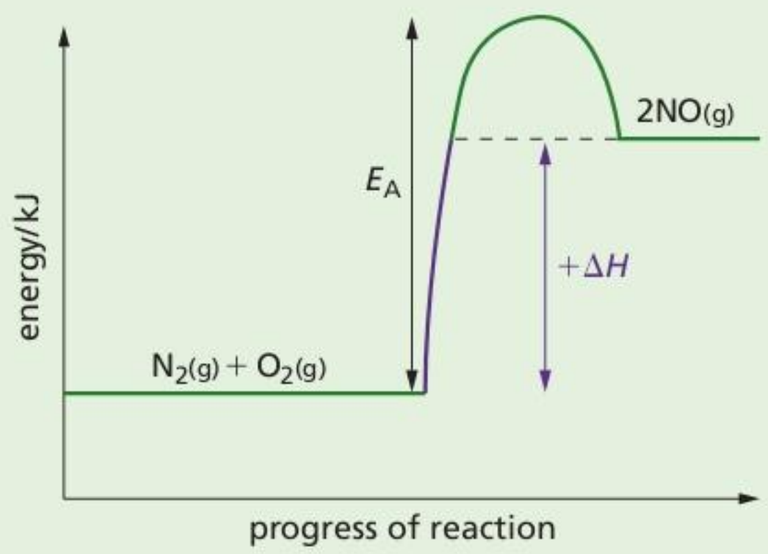


Figure 6.21 Energy level diagram for nitrogen/oxygen.

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 ΔH value in units of kJ mol^{-1}. For example:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$
$$\Delta H = -728\ \text{kJ mol}^{-1}$$

This ΔH 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 ($H^+(aq)$) is neutralised.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l) \quad \Delta H = -57\ \text{kJ mol}^{-1}$$

This process occurs in the titration of an alkali by an acid to produce a neutral solution (Chapter 8, p. 124).

Questions

1 Using the bond energy data given in Table 6.3:
- a Calculate the enthalpy of combustion of ethanol, a fuel added to petrol in some countries.
- b Draw an energy level diagram to represent this combustion process.
- c How does this compare with the enthalpy of combustion of heptane (C_7H_{14}), a major component of petrol, of -4853 kJ mol^{-1}?
- d How much energy is released per gram of ethanol and heptane burned?

2 How much energy is released if:
- a 0.5 mole of methane is burned?
- b 5 moles of methane are burned?
- c 4 g of methane are burned? (A_r: C = 12; H = 1)

3 How much energy is released if:
- a 2 moles of hydrogen ions are neutralised?
- b 0.25 mole of hydrogen ions is neutralised?
- c 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 ΔH_{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 ΔH_{vap}. Figure 6.22 shows an energy level diagram representing both the fusion and the vaporisation processes.

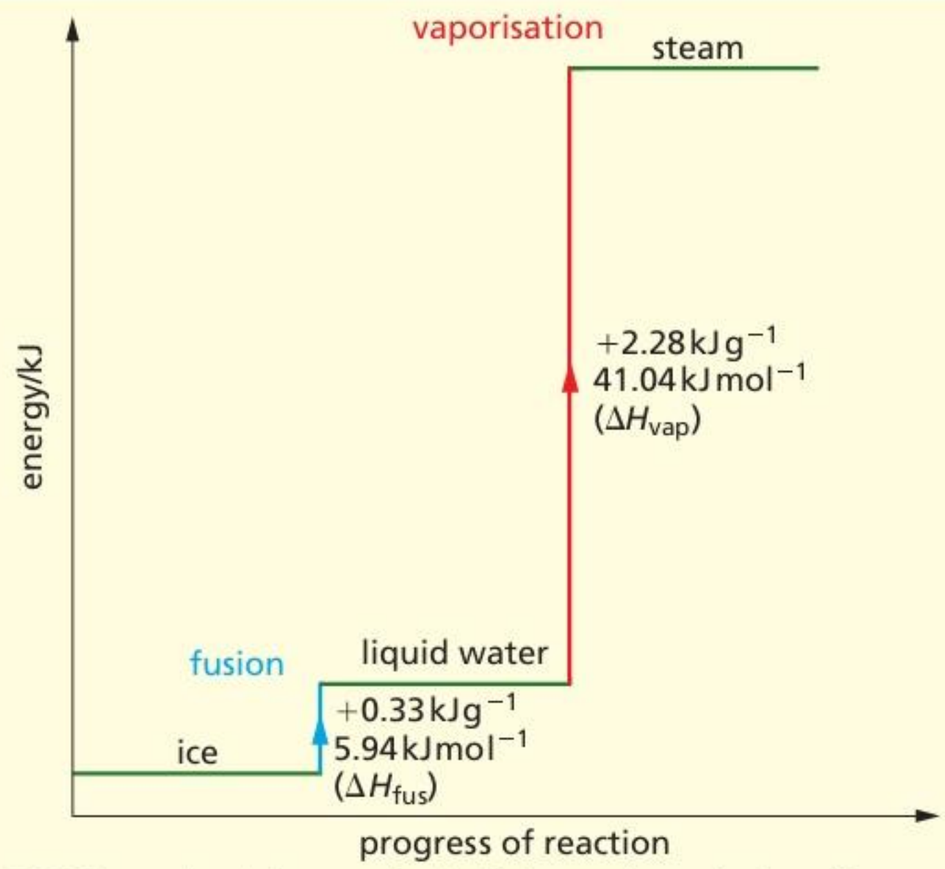


Figure 6.22 Energy level diagram for the fusion and vaporisation of water.

Questions

1 Describe the energy changes which take place when the processes described in this section, with water, are reversed.

2 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 zinc dissolves in the dilute sulfuric acid, producing zinc ions ($Zn^{2+}(aq)$) and releasing two electrons.

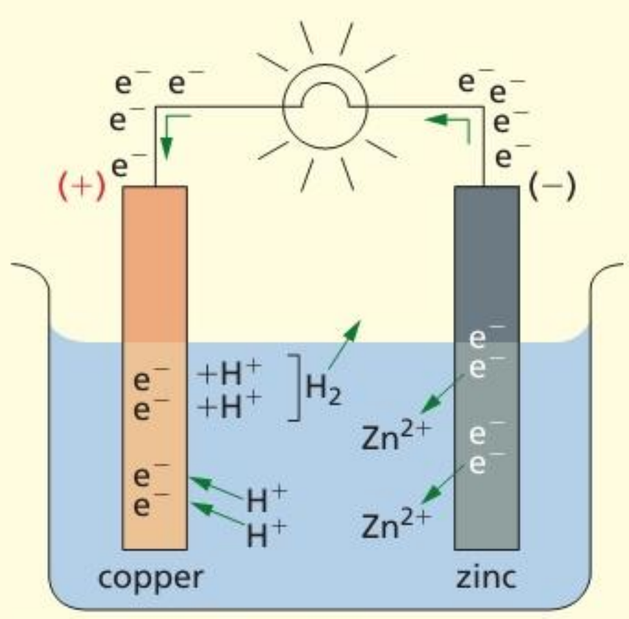
$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$$

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.

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

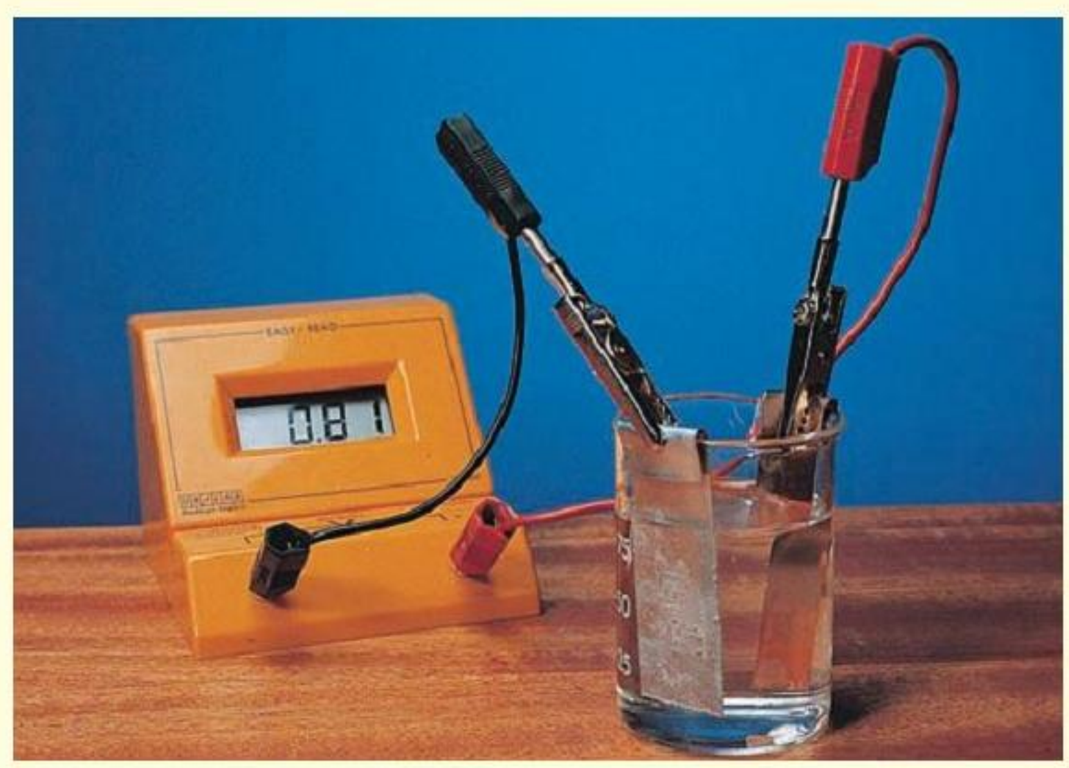
Slowly, the zinc electrode dissolves in the acid and the bulb will then go out. If the zinc is replaced by a more reactive metal, such as magnesium, then the bulb glows more brightly. Magnesium loses electrons more easily as it reacts faster with the dilute acid.

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.



a A simple chemical cell.

Figure 6.23



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:

$$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$$

The OH^- ions formed are removed from the fuel cell by reaction with H_2:

$$H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$$

The electrons produced by this process pass around an external circuit to the cathode.

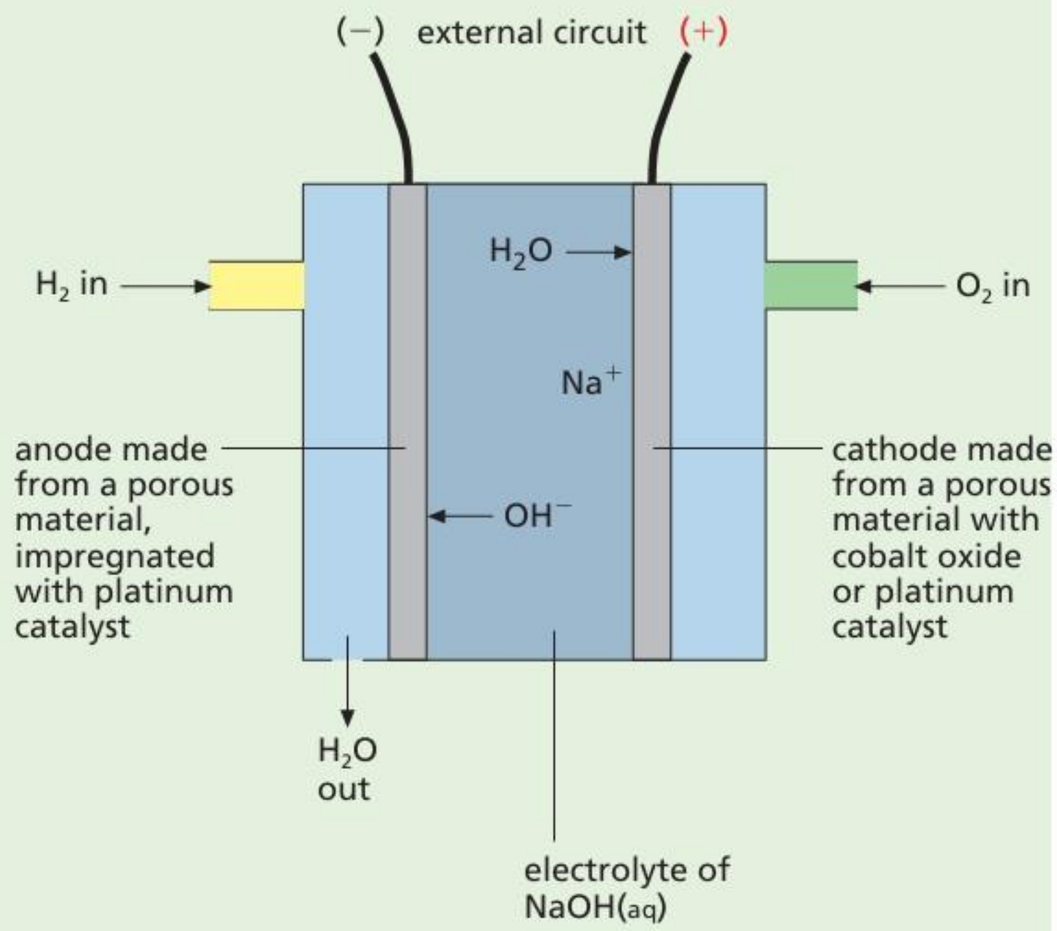


Figure 6.25 A diagrammatic view of a fuel cell.

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 ΔH, 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.

(i) The breaking of a chemical bond.

(ii) The forming of a chemical bond.

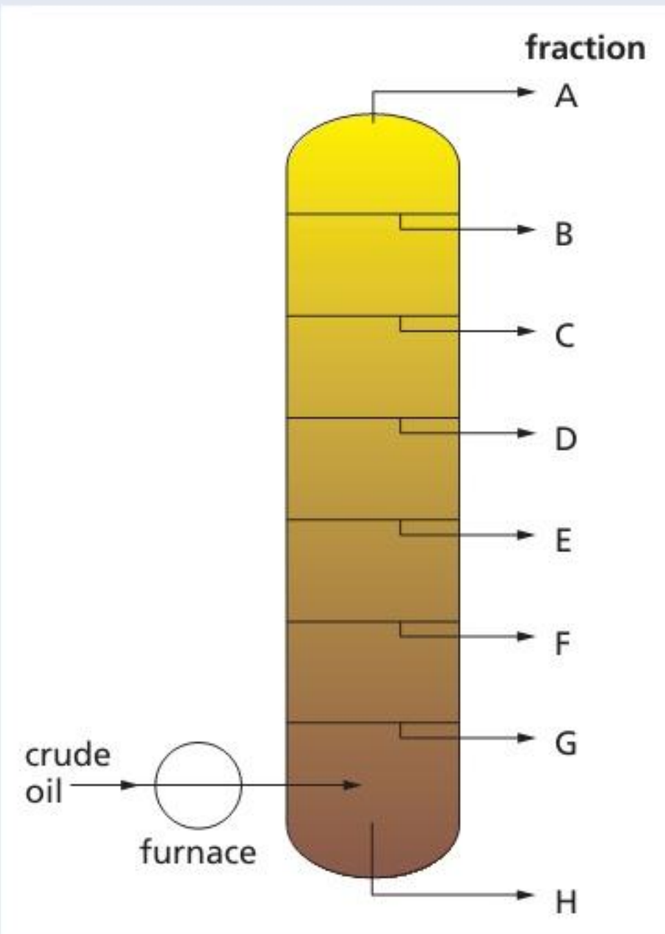
b The table below shows the bond energy data for a series of covalent bonds.

(i) Use the information given in the table to calculate the overall enthalpy change for the combustion of ethanol producing carbon dioxide and water.

(ii) Is the process in **(i)** endothermic or exothermic?

Bond	Bond energy/ $kJ\,mol^{-1}$
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.

a Hydroelectric power is a relatively cheap source of electricity.

b Geothermal energy is a non-polluting form of energy.

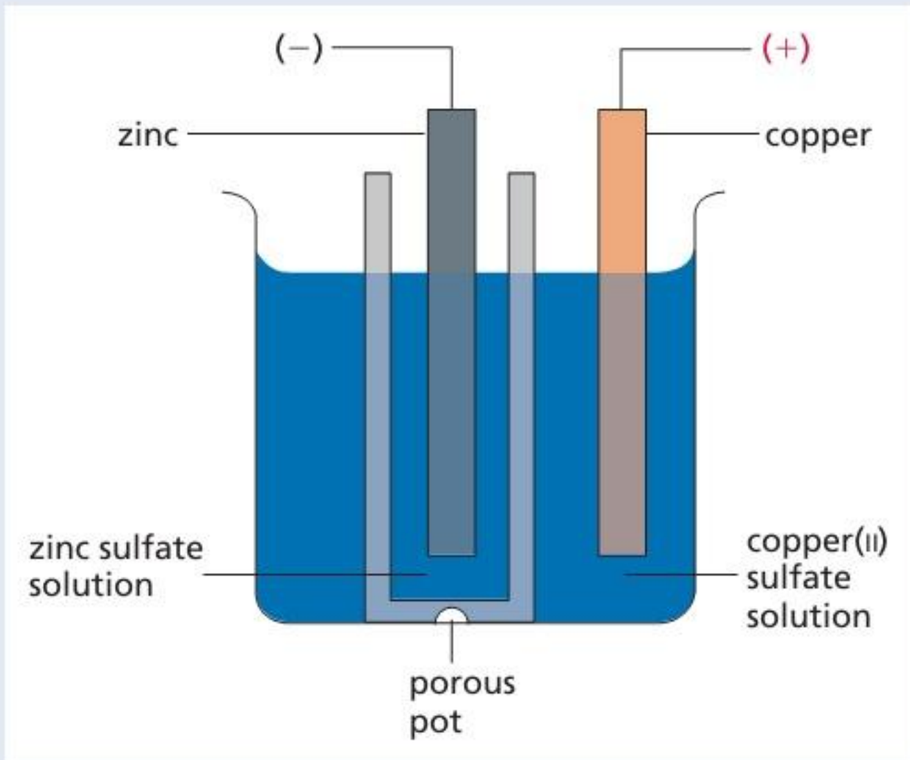
c A disadvantage of wind power is that it causes noise pollution.

d The by-product from the process by which methane is generated by the digestion of animal waste is an excellent fertiliser.

e The fission of uranium-235 in a nuclear reactor is an example of a chain reaction.

f 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:

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

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:

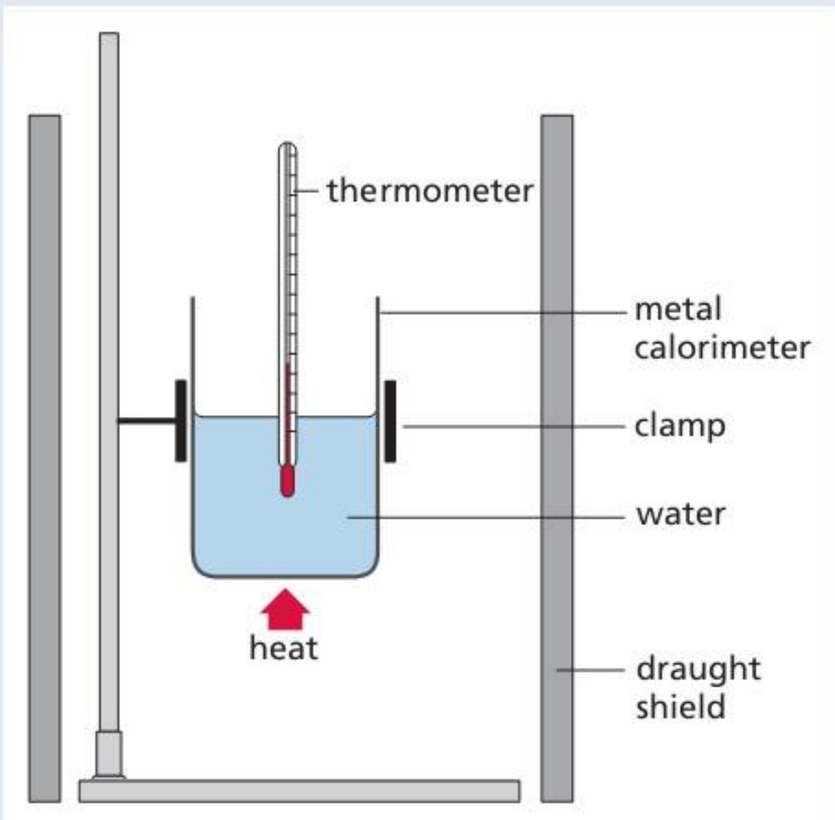
$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$$
$$\Delta H = -57\ kJ\ mol^{-1}$$

(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:

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$
$$\Delta H = +575\ kJ\ mol^{-1}$$

(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^3
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^{-3})

Heat energy given to water = mass of water/g × 4.2 $J\,g^{-1}\,°C^{-1}$ × 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_2H_5OH)?
(A_r: 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^{-1} 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^{-3} hydrochloric acid and 1 mol dm^{-3} 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^3
Volume of acid used = 50 cm^3
Temperature rise = 5 °C
(Density of water = 1 g cm^{-3})

Heat energy given out during reaction = mass of water/g × 4.2 $J\,g^{-1}\,°C^{-1}$ × 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 $-57\,kJ\,mol^{-1}$. 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:
a fuel for a cigarette lighter
b fuel for a camping stove
c fuel for an aeroplane
d fuel for an underground transport system
e fuel for a space craft
f 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:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$$
$$\Delta H = -2220\,kJ\,mol^{-1}$$

a What are the chemical names for 'propagas' and natural gas?
b Would you expect the heat generated per mole of 'propagas' burned to be greater than that for natural gas? Explain your answer.
c What is 'propagas' obtained from?
d Calculate:
(i) the mass of 'propagas' required to produce 5550 kJ of energy
(ii) the heat energy produced by burning 0.5 mole of 'propagas'
(iii) the heat energy produced by burning 11 g of 'propagas'
(iv) the heat energy produced by burning $2000\,dm^3$ of 'propagas'.
(A_r: H = 1; C = 12; O = 16. One mole of any gas occupies $24\,dm^3$ 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.

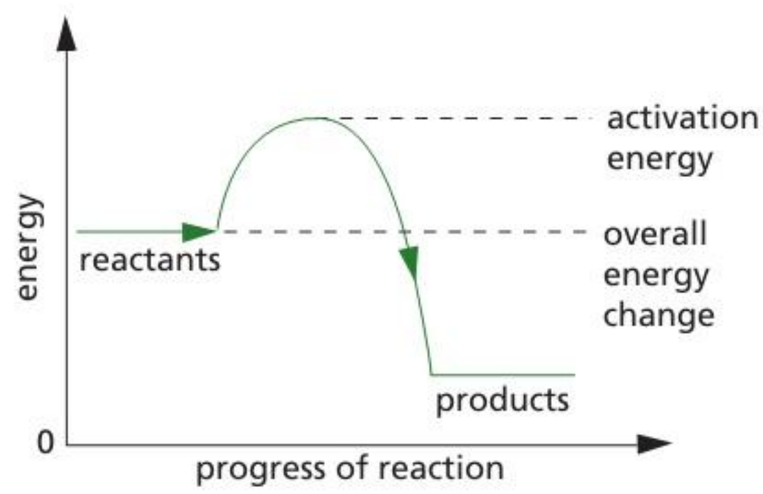


Figure 7.2 Energy level diagram showing activation energy.

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.

hydrochloric acid + calcium carbonate → calcium chloride + carbon dioxide + water

$$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Figure 7.3 The powdered limestone (left) reacts faster with the acid than the limestone in the form of lumps.

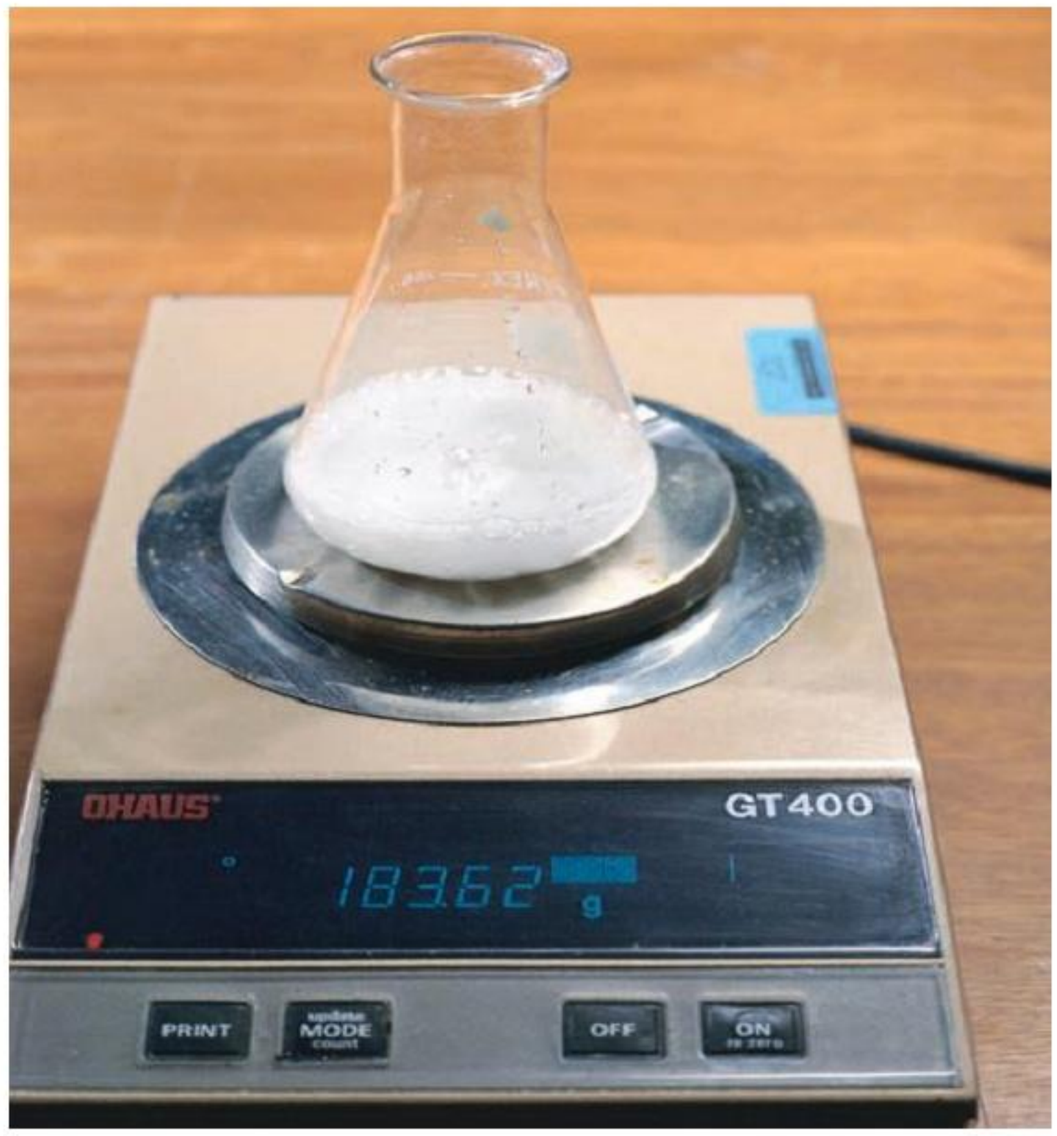
The rates at which the two reactions occur can be found by measuring either:

- 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.



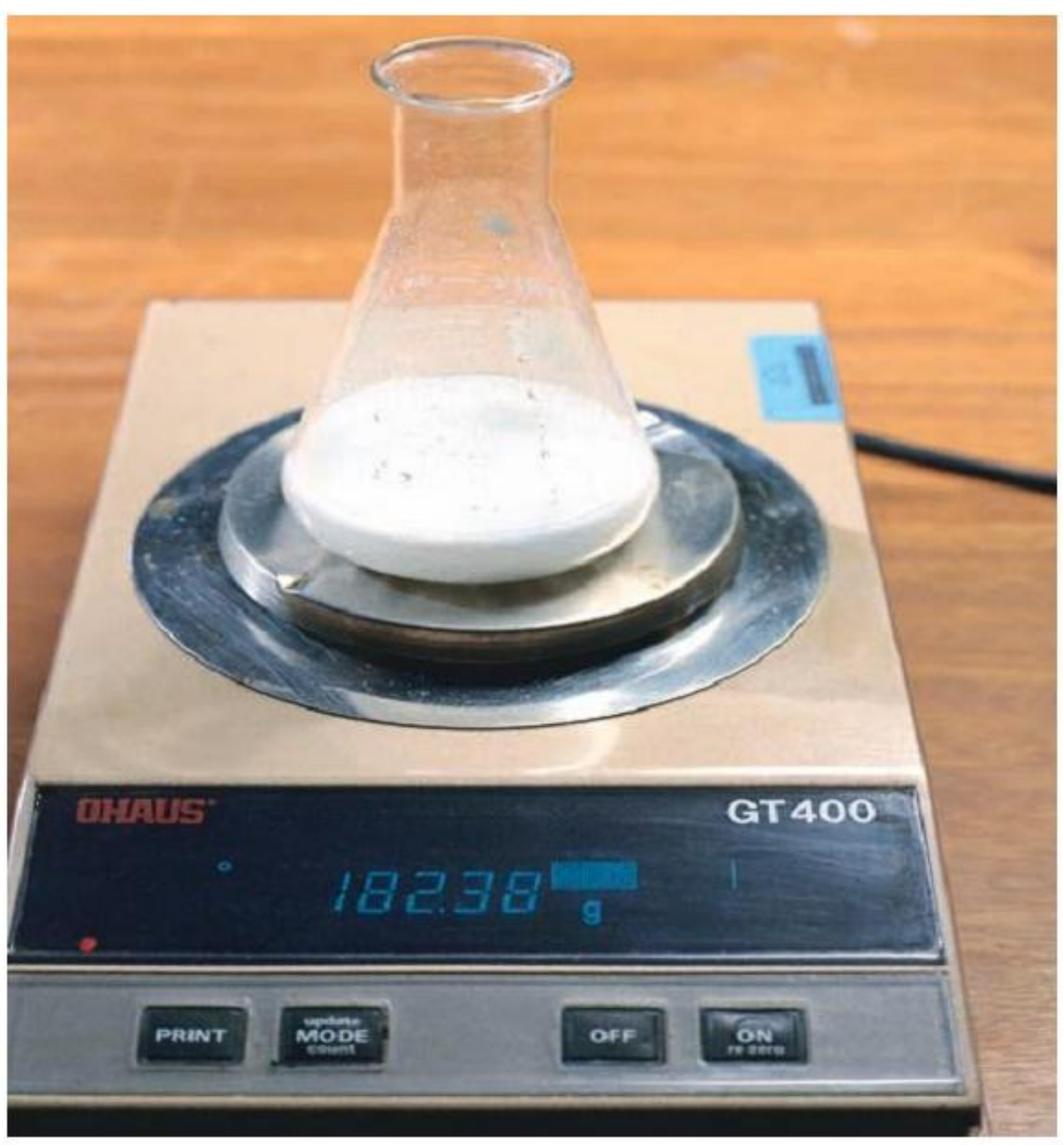


Figure 7.4 After 60 seconds the mass has fallen by 1.24 g.

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.

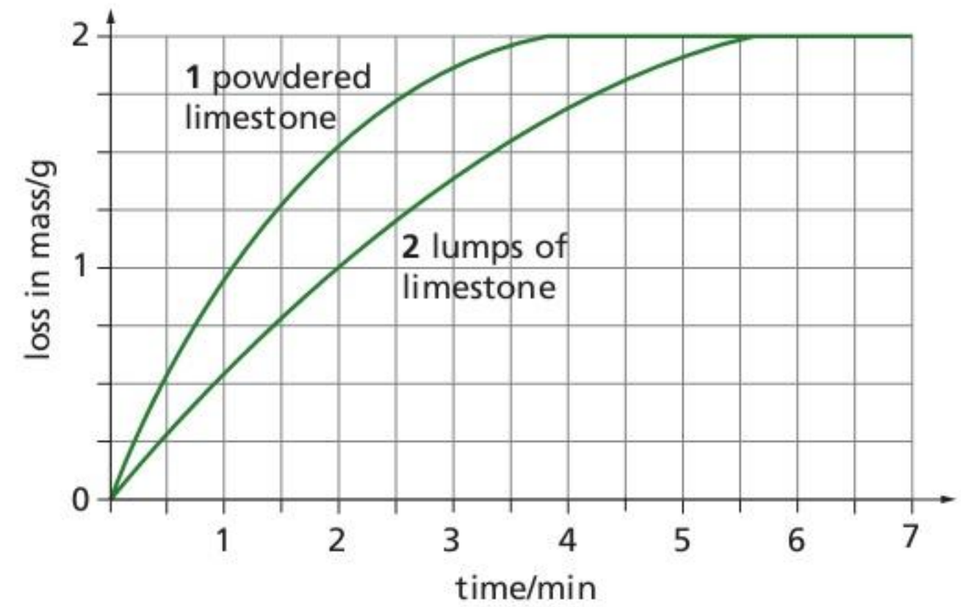


Figure 7.5 Sample results for the limestone/acid experiment.

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.

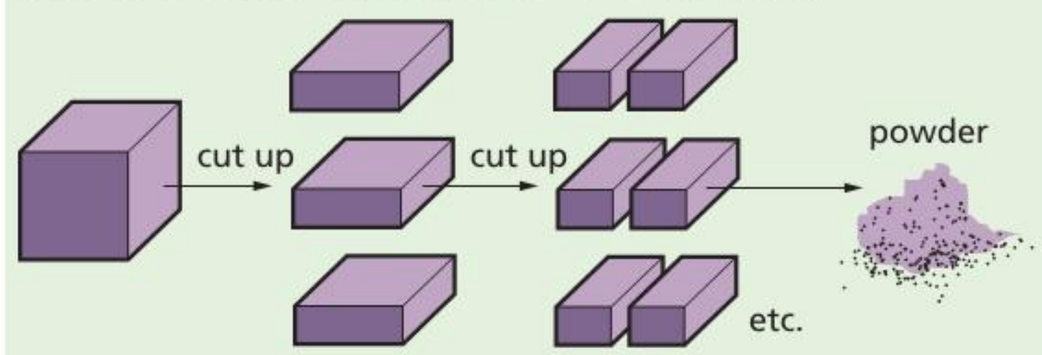


Figure 7.6 A powder has a larger surface area.

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_2 gas/cm^3	0	15	24	28	31	33	35	35	35	35	35

a Plot a graph of the total volume of CO_2 against time.
b At which point is the rate of reaction fastest?
c What volume of CO_2 was produced after 1 minute 15 seconds?
d How long did it take to produce 30 cm^3 of CO_2?

Concentration

A yellow precipitate is produced in the reaction between sodium thiosulfate and hydrochloric acid.

sodium thiosulfate + hydrochloric acid → sodium chloride + sulfur + sulfur dioxide + water

$$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + S(s) + SO_2(g) + H_2O(l)$$

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.

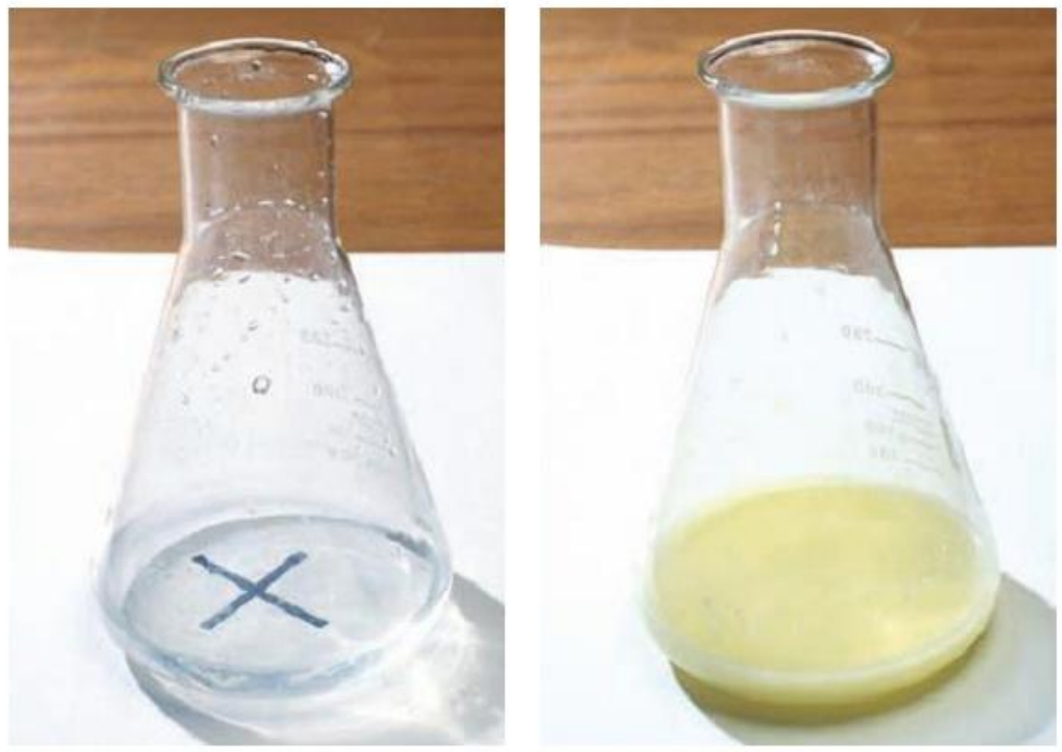


Figure 7.8 The precipitate of sulfur obscures the cross.

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.

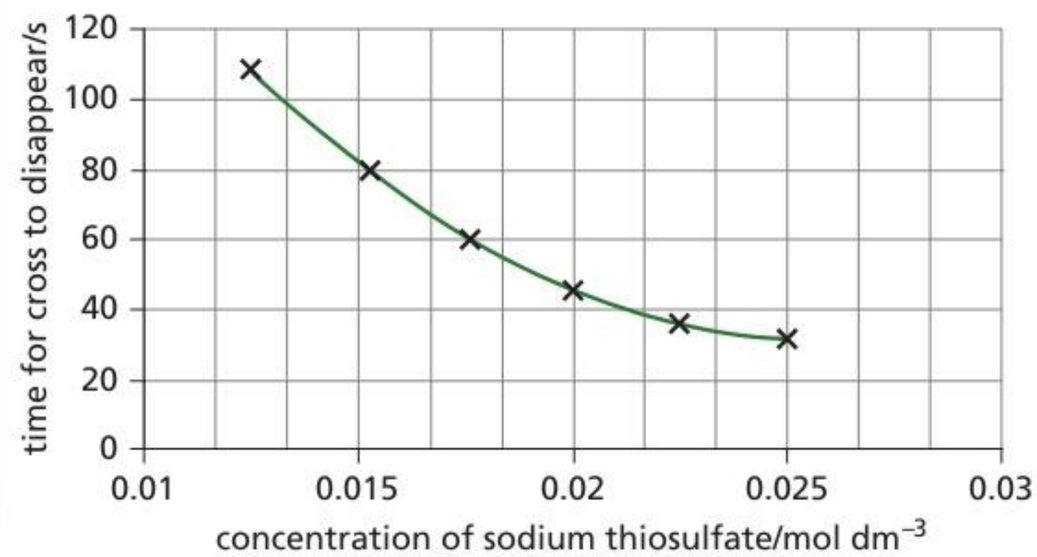


Figure 7.9 Sample data for the sodium thiosulfate/acid experiment at different concentrations of sodium thiosulfate.

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.

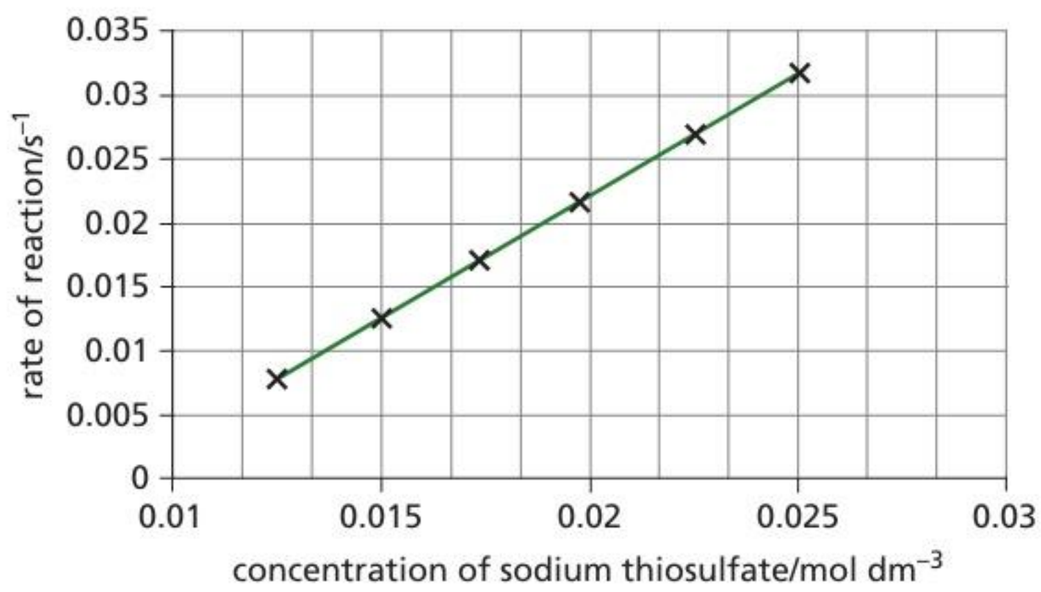


Figure 7.10 Graph to show the rate of reaction against concentration.

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.

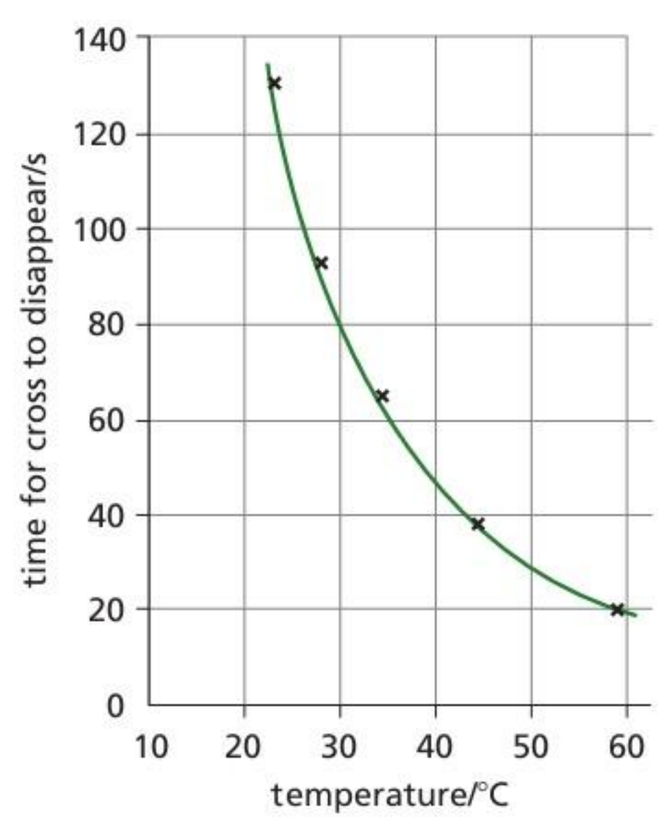


Figure 7.11 Sample data for the sodium thiosulfate/acid experiment at different temperatures.

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_3, 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

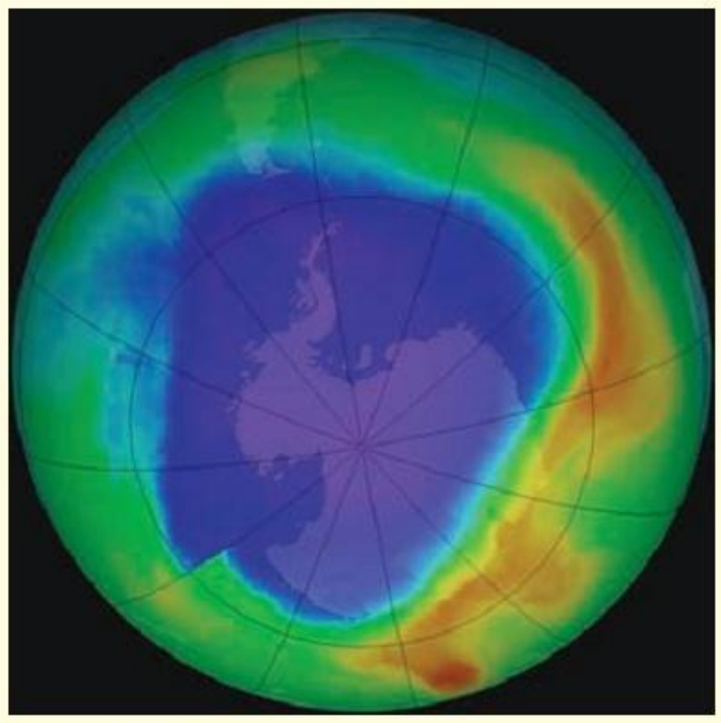


Figure 7.12 Diagram showing the hole in the ozone layer over Antarctica

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:

$$CF_2Cl_2 \xrightarrow{\text{uv light}} CF_2Cl + \underset{\text{chlorine radical}}{Cl}$$

This chlorine radical then reacts with the ozone, destroying it:

$$Cl(g) + O_3(g) \rightarrow OCl(g) + O_2(g)$$

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 (AgBr). The emulsion used is similar for both black-and-white and colour film.

When light hits a silver bromide crystal, silver cations (Ag^+) accept an electron (reduction) from the bromide ions (Br^-), which are oxidised. Hence, silver atoms and bromine atoms are produced in the emulsion.

$$\text{silver ion} + \text{electron} \rightarrow \text{silver atom}$$
$$Ag^+ + e^- \rightarrow Ag$$

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.

$$\text{hydrogen peroxide} \rightarrow \text{water} + \text{oxygen}$$
$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

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.

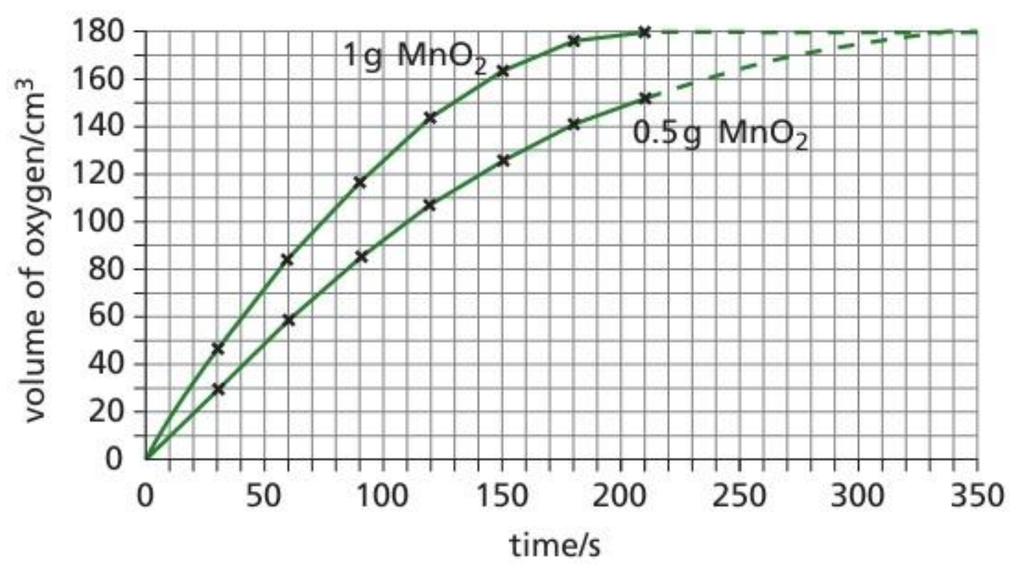


Figure 7.13 Sample data for differing amounts of MnO_2 catalyst.

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).

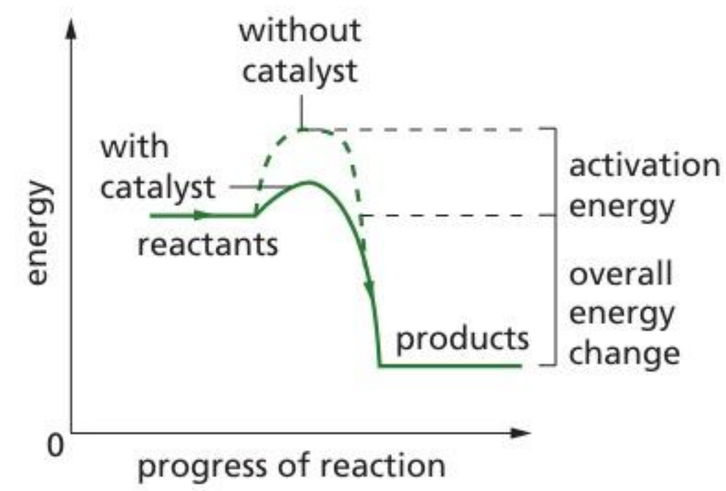


Figure 7.14 Energy level diagram showing activation energy, with and without a catalyst.

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.

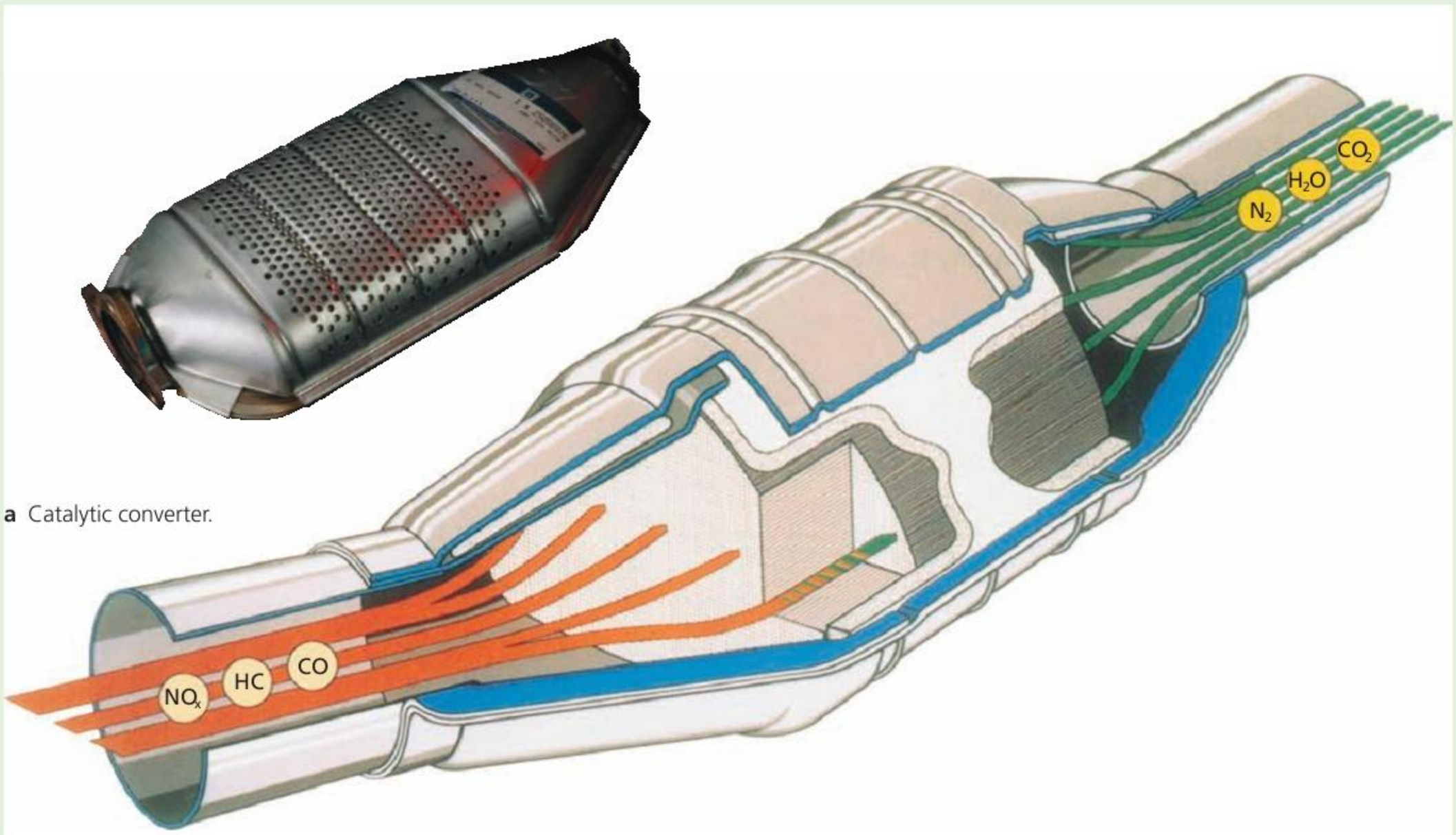
carbon monoxide + oxygen → carbon dioxide

$$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$$

nitrogen(II) oxide + carbon monoxide → nitrogen + carbon dioxide

$$2NO(g) + 2CO(g) \rightarrow N_2(g) + 2CO_2(g)$$

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



a Catalytic converter.

b A section through a catalytic converter.

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 on 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:

$$\text{hydrogen peroxide} \xrightarrow{\text{catalase}} \text{water} + \text{oxygen}$$

$$2H_2O_2(aq) \xrightarrow{\text{catalase}} 2H_2O(l) + O_2(g)$$

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 A car exhaust pipe will rust much faster if the car is in constant use.
   b Vegetables cook faster when they are chopped up.
   c Industrial processes become more economically viable if a catalyst can be found for the reactions involved.
   d In fireworks it is usual for the ingredients to be powdered.
   e Tomatoes ripen faster in a greenhouse.
   f 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^3	0	29	55	79	98	118	133	146
Volume for 0.5 g/cm^3	0	45	84	118	145	162	174	182

   a Draw a diagram of the apparatus you could use to carry out these experiments.
   b Plot a graph of the results.
   c Is the manganese(IV) oxide acting as a catalyst in this reaction? Explain your answer.
   d (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).
   e Why does the slope of the graph become less steep as the reaction proceeds?
   f What volume of gas has been produced when using 0.3 g of manganese(IV) oxide after 50 s?
   g How long did it take for 60 cm^3 of gas to be produced when the experiment was carried out using 0.5 g of the manganese(IV) oxide?
   h 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?
      (i) zinc granules + dilute nitric acid
      (ii) 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

   a Plot the results of the experiment.
   b Which of the results would appear to be incorrect? Explain your answer.
   c Write a balanced chemical equation to represent the reaction taking place.
   d Why did the mass of the flask and its contents decrease?
   e Why was the plug of cotton wool used?
   f How does the rate of reaction change during this reaction? Explain this using particle theory.
   g How long did the reaction last?
   h How long did it take for half of the reaction to occur?

5 a What is a catalyst?
   b List the properties of catalysts.
   c Name the catalyst used in the following processes:
      (i) the Contact process
      (ii) the Haber process
      (iii) the hydrogenation of unsaturated fats.
   d Which series of metallic elements in the Periodic Table (p. 136) do the catalysts you have named in **c** belong to?
   e 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:

$$CuCO_3(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + CO_2(aq) + H_2O(l)$$

a 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^3 of gas.

b Find on your graph the part which shows:
(i) where the reaction is at its fastest
(ii) when the reaction has stopped.

c Calculate the mass of copper(II) carbonate used to produce 60 cm^3 of carbon dioxide. (A_r: C = 12; O = 16; Cu = 63.5. One mole of a gas occupies 24 dm^3 at room temperature and pressure (rtp).)

d Sketch a further graph using the same axes to show what happens to the rate at which the gas is produced if:
(i) the concentration of the acid is decreased
(ii) 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.

a Why are these regulations necessary?

b Which gases are removed by catalytic converters?

c Which metals are often used as catalysts in catalytic converters?

d What does the term 'poisoned' mean with respect to catalysts?

e The latest converters will also remove unburnt petrol. An equation for this type of reaction is:

$$2C_7H_{14}(g) + 21O_2(g) \rightarrow 14CO_2(g) + 14H_2O(g)$$

(i) Calculate the mass of carbon dioxide produced by 1.96 g of unburnt fuel.
(ii) Convert this mass of carbon dioxide into a volume measured at rtp.
(iii) 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_r: H = 1; C = 12; O = 16. One mole of any gas occupies 24 dm^3 at rtp.)

8 Suggest practical methods by which the rate of reaction can be investigated in each of the following cases:

a magnesium reacting with hydrochloric acid

b 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.

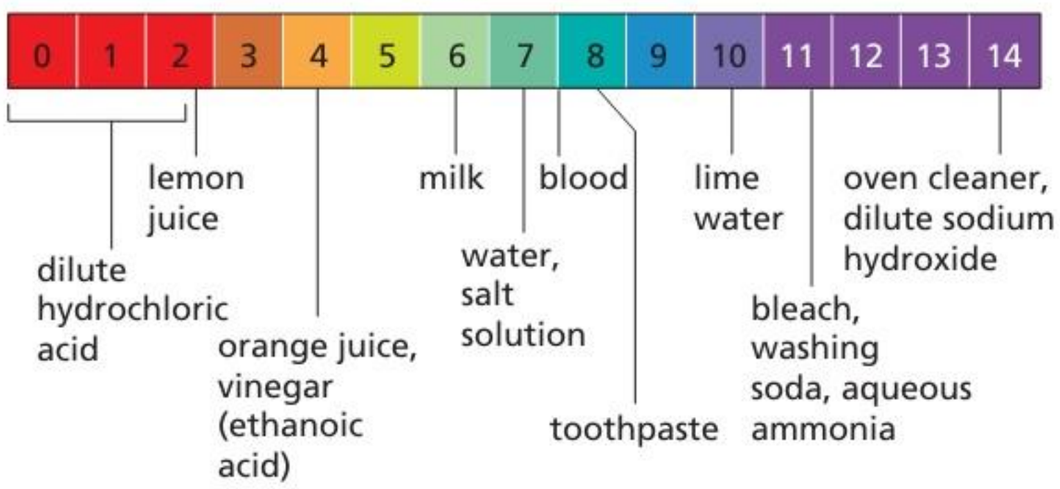
Methyl orange, a common indicator used in titrations (see p. 129) is pink in an acid solution but changes to show a yellow colour in an alkaline solution. Some other indicators are shown in Table 8.1, along with the colours they turn in acids and alkalis.

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 Sørenson. 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.

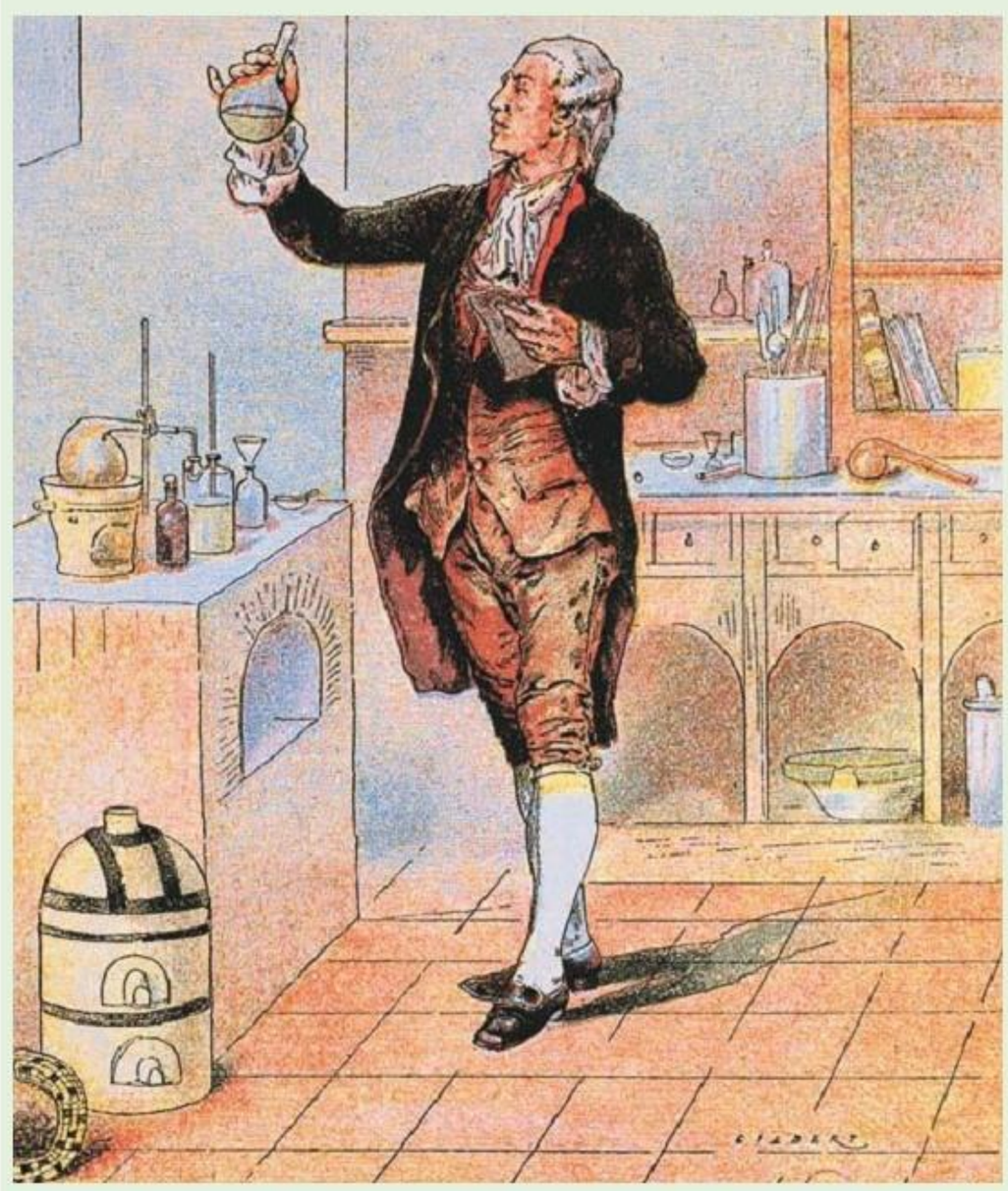


Figure 8.6 Antoine Lavoisier (1714–1794) at work in his laboratory.

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.

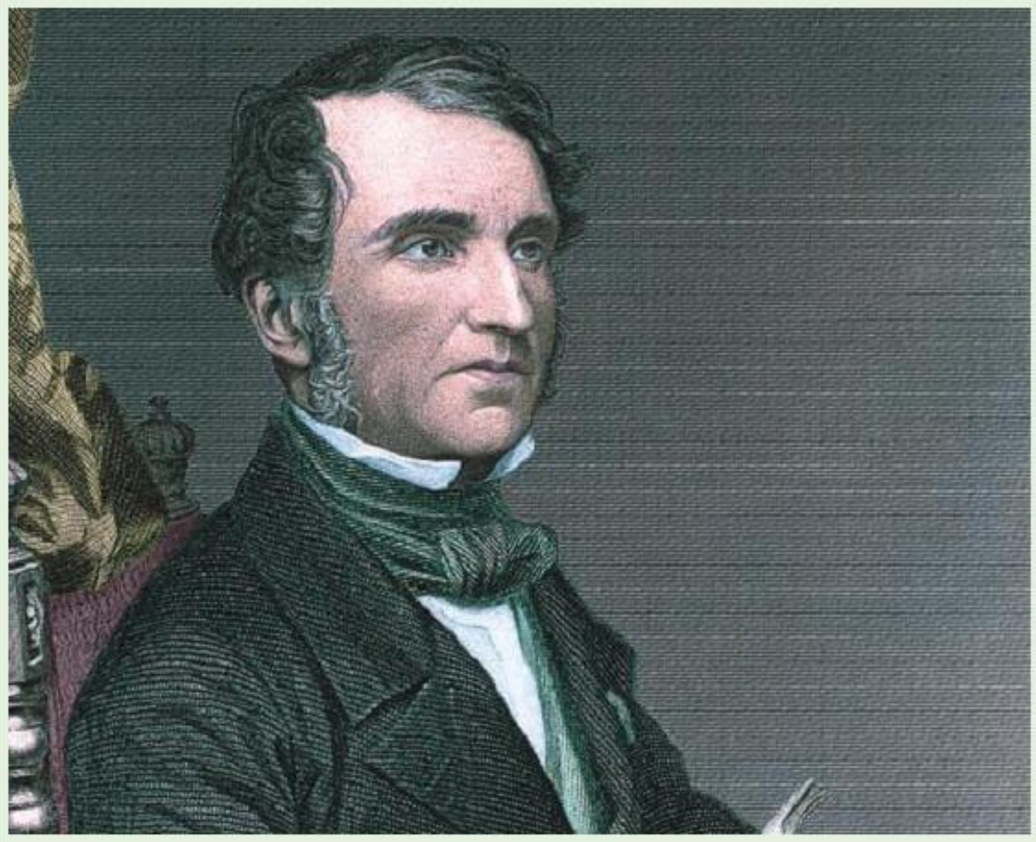


Figure 8.7 Justus von Liebig (1803–1873), who defined acids more closely.

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 (H^+) when they dissolve in water, whereas bases produce hydroxide ions (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, HCl(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 H^+ ion (or proton) donor
- a base as an 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 H^+ ion donor it must have another substance present to accept the H^+ ion. So the water, in the aqueous acid solution, is behaving as a Brønsted–Lowry base and accepting an H^+ ion. Generally:

$$\underset{\text{acid}}{HA(aq)} + \underset{\text{base}}{H_2O(l)} \rightarrow H_3O^+(aq) + A^-(aq)$$

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 OH^-:

$$\underset{\text{base}}{NH_3(aq)} + \underset{\text{acid}}{H_2O(l)} \rightarrow NH_4^+(aq) + OH^-(aq)$$

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).

$$\underset{\text{acid}}{HCl(g)} + \underset{\text{base}}{NH_3(g)} \rightarrow NH_4^+Cl^-(s)$$

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.

hydrogen chloride $\xrightarrow{\text{water}}$ hydrogen ions + chloride ions

$$HCl(g) \xrightarrow{\text{water}} H^+(aq) + Cl^-(aq)$$

For hydrochloric acid *all* the hydrogen chloride molecules break up to form H^+ ions and 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 ($H^+(aq)$) 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.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

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.

ethanoic acid $\overset{\text{water}}{\rightleftharpoons}$ hydrogen ions + ethanoate ions

$$CH_3COOH(l) \overset{\text{water}}{\rightleftharpoons} H^+(aq) + CH_3COO^-(aq)$$

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, $H^+(aq)$. 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, $OH^-(aq)$, when dissolved in water. Sodium hydroxide is a **strong alkali** because when it dissolves in water its lattice breaks up completely to produce ions.

$$\text{sodium hydroxide} \xrightarrow{\text{water}} \text{sodium ions} + \text{hydroxide ions}$$

$$NaOH(s) \longrightarrow Na^+(aq) + OH^-(aq)$$

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.

$$\text{ammonia} + \text{water} \rightleftharpoons \text{ammonium ions} + \text{hydroxide ions}$$

$$NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

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.

hydrochloric acid	+	sodium hydroxide	→	sodium chloride	+	water
$HCl(aq)$	+	$NaOH(aq)$	→	$NaCl(aq)$	+	$H_2O(l)$

As we have shown, when both hydrochloric acid and sodium hydroxide dissolve in water the ions separate completely. We may therefore write:

$$H^+(aq)Cl^-(aq) + Na^+(aq)OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$

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:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

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 ($H^+(aq)$) reacting with the ion which causes alkalinity ($OH^-(aq)$) to produce neutral water ($H_2O(l)$).

Questions

1 Complete the following equations:
   a $CH_3COOH + NaOH \rightarrow$
   b $H_2SO_4 + KOH \rightarrow$
   c $NH_3 + HBr \rightarrow$
   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 (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 (Na_2CO_3)
Ethanoic acid	Ethanoates	Sodium ethanoate (CH_3COONa)
Hydrochloric acid	Chlorides	Potassium chloride (KCl)
Nitric acid	Nitrates	Potassium nitrate (KNO_3)
Sulfuric acid	Sulfates	Sodium sulfate (Na_2SO_4)

Methods of preparing soluble salts

There are four general methods of preparing soluble salts:

Acid + metal

This method can only be used with the less reactive metals. It would be very dangerous to use a reactive metal such as sodium in this type of reaction. The metals usually used in this method of salt preparation are the **MAZIT** metals, that is, **m**agnesium, **a**luminium, **z**inc, **i**ron and **t**in. A typical experimental method is given below.

Excess magnesium ribbon is added to dilute nitric acid. During this addition an effervescence is observed due to the production of hydrogen gas. In this reaction the hydrogen ions from the nitric acid gain electrons from the metal atoms as the reaction proceeds.

hydrogen ions + electrons (from metal) → hydrogen gas

$$2H^+ + 2e^- \rightarrow H_2(g)$$

How would you test the gas to show that it was hydrogen? What would be the name and formula of the compound produced during the test you suggested?

magnesium + nitric acid → magnesium nitrate + hydrogen

$$Mg(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$$

Figure 8.11 The excess magnesium is filtered in this way.

The excess magnesium is removed by filtration (Figure 8.11).

The magnesium nitrate solution is evaporated slowly to form a saturated solution of the salt (Figure 8.12).

The hot concentrated magnesium nitrate solution produced is tested by dipping a cold glass rod into it. If salt crystals form at the end of the rod the solution is ready to crystallise and is left to cool. Any crystals produced on cooling are filtered and dried between clean tissues.

Figure 8.12 The solution of magnesium nitrate is concentrated by slow evaporation.

Acid + carbonate

This method can be used with any metal carbonate and any acid, providing the salt produced is soluble. The typical experimental procedure is similar to that carried out for an acid and a metal. For example, copper(II) carbonate would be added in excess to dilute nitric acid. Effervescence would be observed due to the production of carbon dioxide.

How would you test the gas to show it was carbon dioxide? Write an equation to help you explain what is happening during the test you have chosen.

copper(II) carbonate + nitric acid → copper(II) nitrate + carbon dioxide + water

$$CuCO_3(s) + 2HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + CO_2(g) + H_2O(l)$$

Metal carbonates contain carbonate ions, CO_3^{2-}. In this reaction the carbonate ions react with the hydrogen ions in the acid.

carbonate ions	+	hydrogen ions	→	carbon dioxide	+	water
$CO_3^{2-}(aq)$	+	$2H^+(aq)$	→	$CO_2(g)$	+	$H_2O(l)$

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 $OH^-(aq)$ 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 (Fe_2O_3)
Potassium hydroxide (KOH)	Copper(II) oxide (CuO)
Calcium hydroxide ($Ca(OH)_2$)	Lead(II) oxide (PbO)
Ammonia solution ($NH_3(aq)$)	Magnesium oxide (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,

hydrochloric acid	+	sodium hydroxide	→	sodium chloride	+	water
$HCl(aq)$	+	$NaOH(aq)$	→	$NaCl(aq)$	+	$H_2O(l)$

As previously discussed on p. 116, this reaction can best be described by the ionic equation:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

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.

sulfuric acid	+ copper(II) oxide	→	copper(II) sulfate	+	water
$H_2SO_4(aq)$	+ $CuO(s)$	→	$CuSO_4(aq)$	+	$H_2O(l)$

Metal oxides contain the oxide ion, O^{2-}. The ionic equation for this reaction is therefore:

$$2H^+(aq) + O^{2-}(s) \rightarrow H_2O(l)$$

or

$$CuO(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l)$$

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:

barium chloride	+ sodium sulfate	→ barium sulfate	+ sodium chloride
$BaCl_2(aq)$	+ $Na_2SO_4(aq)$	→ $BaSO_4(s)$	+ $2NaCl(aq)$

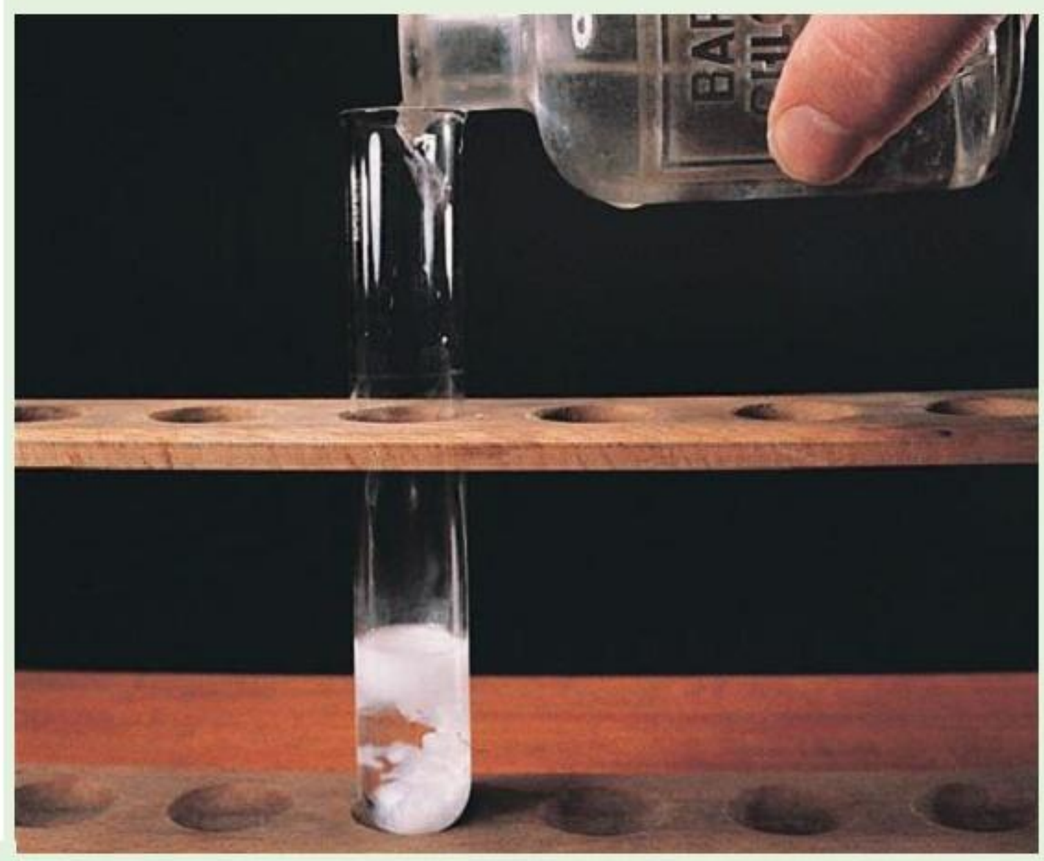


Figure 8.16 When barium chloride solution is added to sodium sulfate a white precipitate of barium sulfate forms.

The ionic equation for this reaction is:

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

This method is sometimes known as **double decomposition** and may be summarised as follows:

soluble salt (AX)	+ soluble salt (BY)	→ insoluble salt (BX)	+ soluble salt (AY)

It should be noted that even salts like barium sulfate dissolve to a very small extent. For example, 1 litre of water will dissolve 2.2×10^{-3} g of barium sulfate at 25 °C. This substance and substances like it are said to be **sparingly 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_2CO_3)	Hydrogencarbonate	Sodium hydrogencarbonate ($NaHCO_3$)
Sulfuric acid (H_2SO_4)	Hydrogensulfate	Potassium hydrogensulfate ($KHSO_4$)

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_4^{2-})

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.

barium ion + sulfate ion → barium sulfate

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

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_3^{2-} ion).

Testing for a chloride (Cl^-), a bromide (Br^-) or an iodide (I^-)

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.

chloride ion + silver ion → silver chloride

$$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$$

If left to stand, the precipitate goes grey (Figure 8.18).

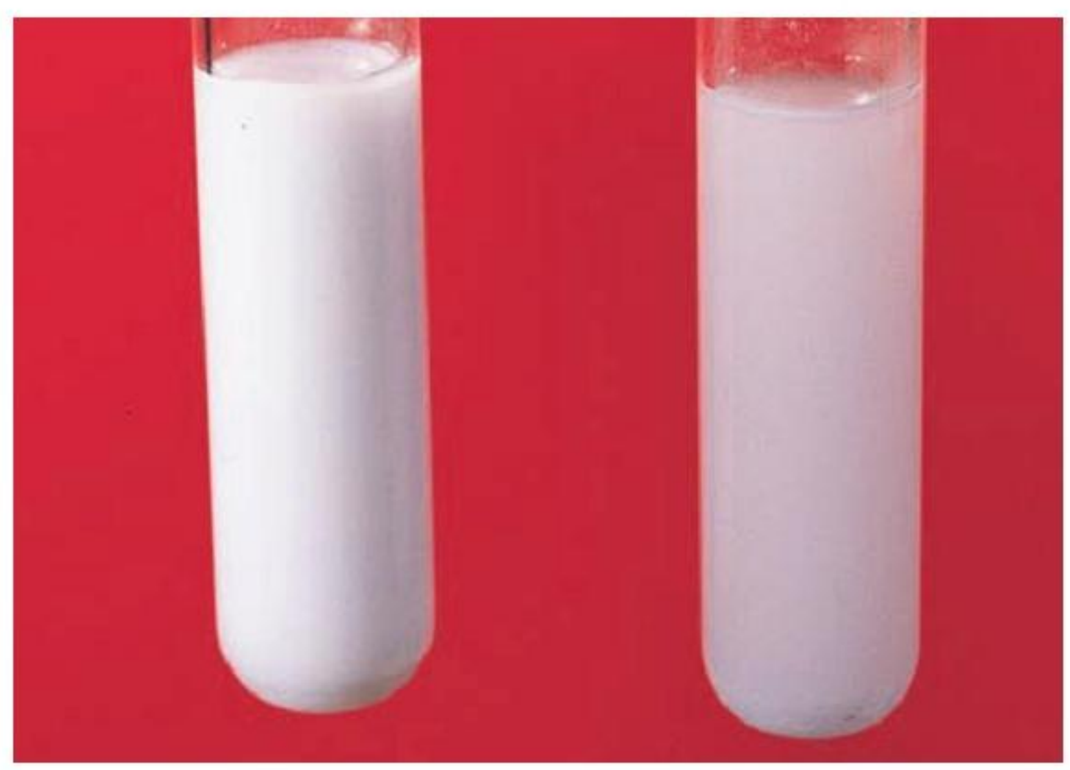


Figure 8.18 If left to stand the white precipitate of silver chloride goes grey. This photochemical change plays an essential part in black and white photography.

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_2.

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).

carbonate ions + hydrogen ions → carbon dioxide + water

$$CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$$

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.

$$3NO_3^-(aq) + 8Al(s) + 5OH^-(aq) + 18H_2O(l) \rightarrow 3NH_3(g) + 8[Al(OH)_4]^-(aq)$$

In the reaction the nitrate ion is reduced, as oxygen is removed from the nitrogen atom, and it gains hydrogen to form ammonia, NH_3. The gain of hydrogen is also a definition of reduction.

Questions

1 Complete the word equations and write balanced chemical equations for the following soluble salt preparations:
   a magnesium + sulfuric acid →
   b calcium carbonate + hydrochloric acid →
   c zinc oxide + hydrochloric acid →
   d potassium hydroxide + nitric acid →
   Also write ionic equations for each of the reactions.

2 Lead carbonate and lead iodide are insoluble. Which two soluble salts could you use in the preparation of each substance? Write
   a a word equation
   b a symbol equation
   c an ionic equation
   to represent the reactions taking place.

3 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.
   a Describe how she could confirm the presence of these three types of salt by simple chemical tests.
   b 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	$CoCl_2.6H_2O$
Copper(II) sulfate pentahydrate	$CuSO_4.5H_2O$
Iron(II) sulfate heptahydrate	$FeSO_4.7H_2O$
Magnesium sulfate heptahydrate	$MgSO_4.7H_2O$
Sodium carbonate decahydrate	$Na_2CO_3.10H_2O$
Sodium hydrogensulfate monohydrate	$NaHSO_4.H_2O$
Sodium sulfate decahydrate	$Na_2SO_4.10H_2O$

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:

copper(II) sulfate pentahydrate → anhydrous copper(II) sulfate + water

$$CuSO_4.5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(g)$$

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.

$$CuSO_4(s) + 5H_2O(l) \rightarrow CuSO_4.5H_2O(s)$$

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:

$$CuSO_4(s) + 5H_2O(l) \rightleftharpoons CuSO_4.5H_2O(s)$$

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**.

$$Na_2CO_3.10H_2O(s) \rightarrow Na_2CO_3.H_2O(s) + 9H_2O(g)$$

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_2O' is treated as an element in the calculation.

Example

Calculate the percentage by mass of water in the salt hydrate $MgSO_4.7H_2O$. (A_r: H = 1; O = 16; Mg = 24; S = 32)

M_r for $MgSO_4.7H_2O$
$= 24 + 32 + (4 \times 16) + (7 \times 18)$
$= 246$

The mass of water as a fraction of the total mass of hydrate

$$= \frac{126}{246}$$

The percentage of water present

$$= \frac{126}{246} \times 100$$

$= 51.2\%$

Questions

1 Calculate the percentage by mass of water in the following salt hydrates:
   a $CuSO_4.5H_2O$
   b $Na_2CO_3.10H_2O$
   c $Na_2S_2O_3.5H_2O$.
   (A_r: H = 1; O = 16; Na = 23; S = 32; Cu = 63.5)
2 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 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 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.

1 25 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.

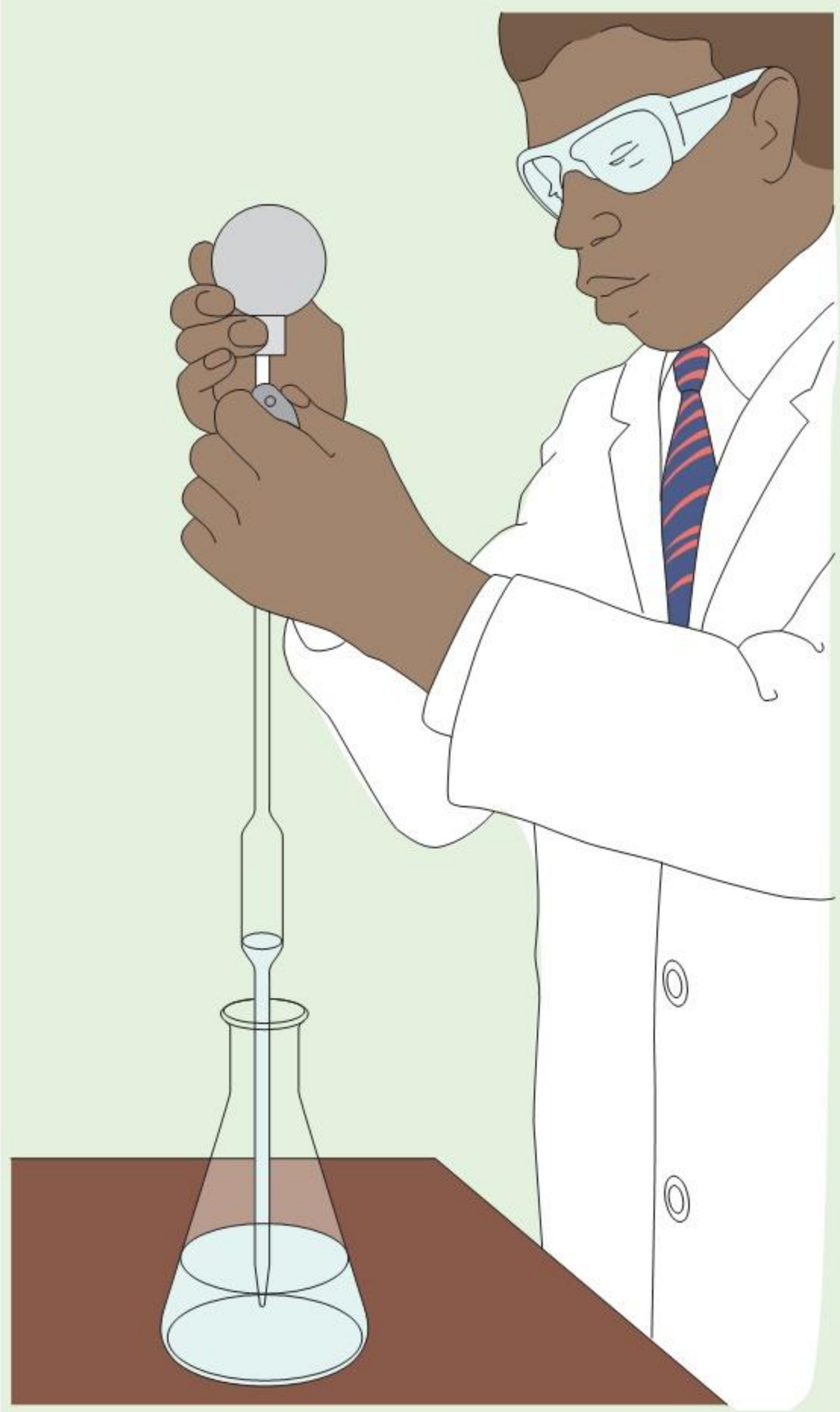


Figure 8.23 Exactly 25.0 cm^3 of sodium hydroxide solution is pipetted into a conical flask.

2 A 0.10 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).

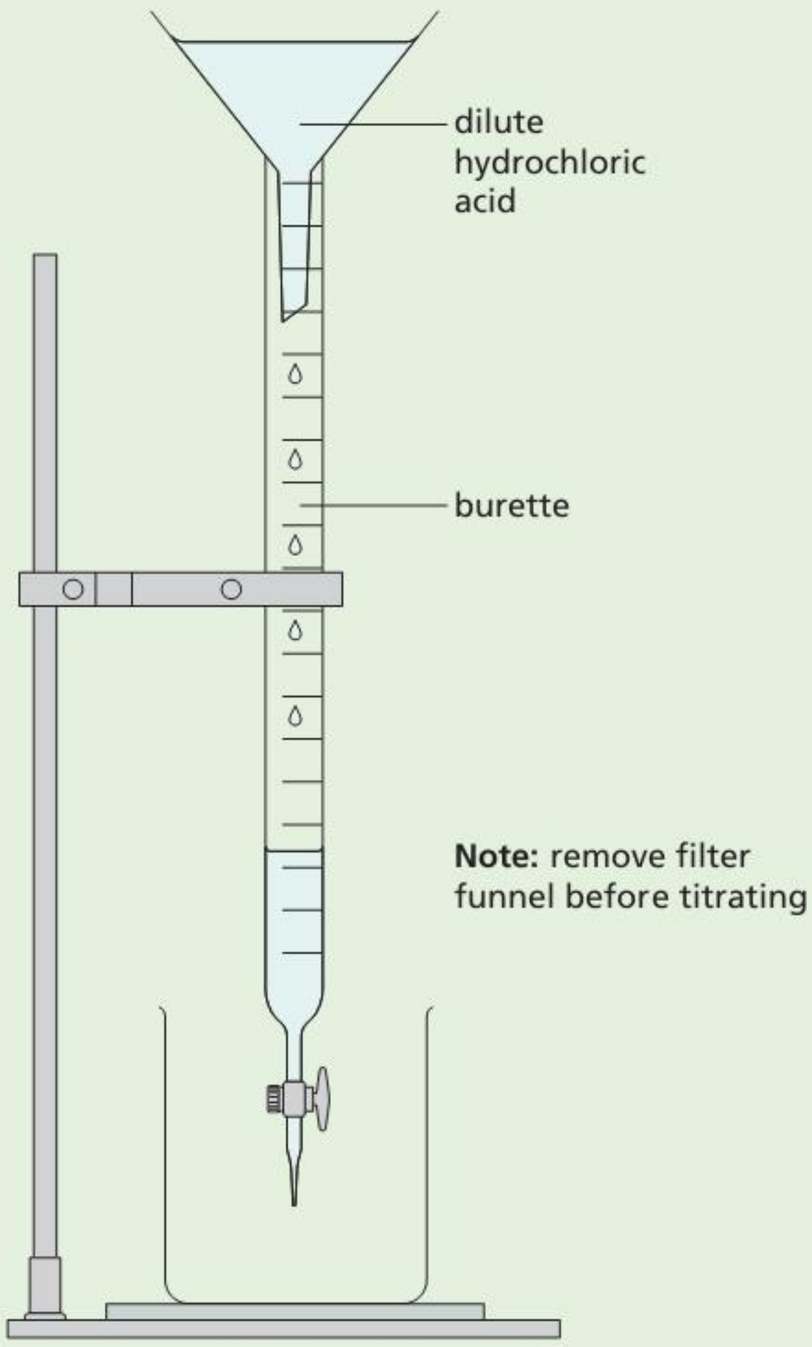


Figure 8.24 The burette is filled up to the zero mark with a 0.10 mol dm^{-3} solution of hydrochloric acid.

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 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.

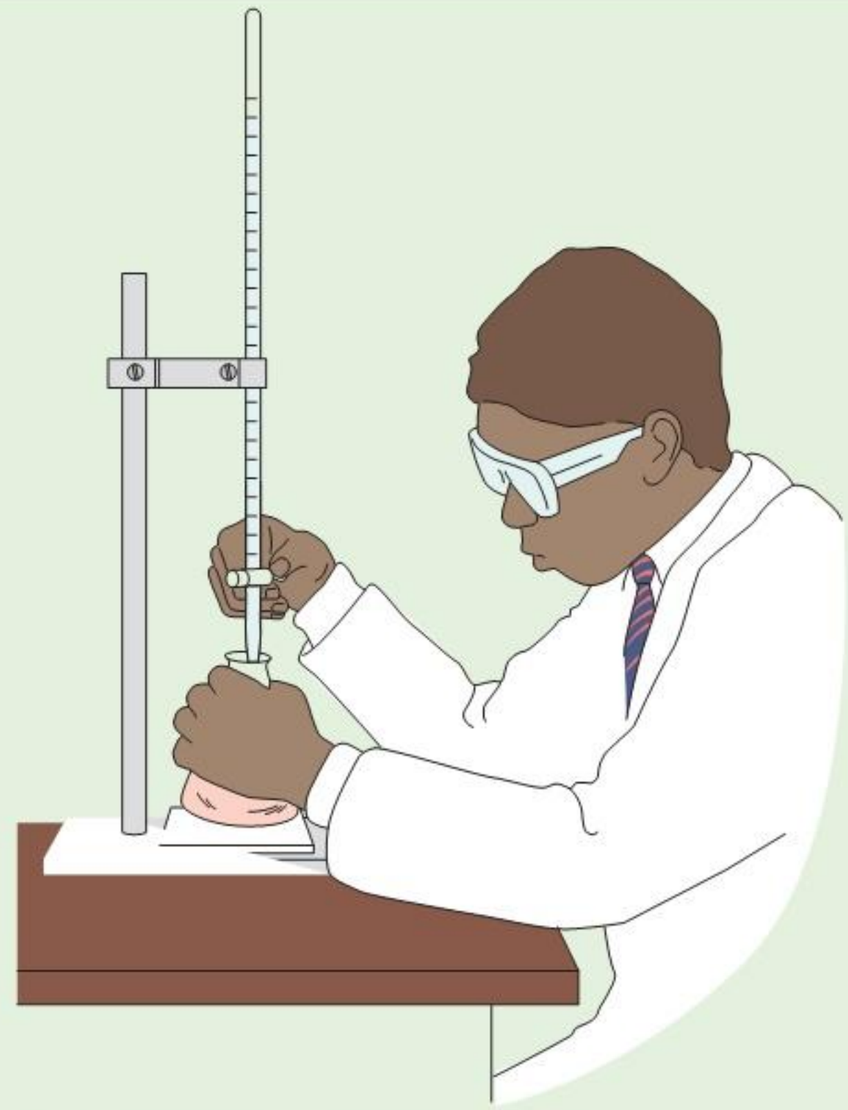


Figure 8.25 The titration is carried out accurately.

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 cm^3 of each other). Some sample data are shown below.

Volume of sodium hydroxide solution
= 25.0 cm^3

Average volume of 0.10 mol dm^{-3} solution of hydrochloric acid added
= 21.0 cm^3

The neutralisation reaction which has taken place is:

hydrochloric acid	+	sodium hydroxide	→	sodium chloride	+	water
$HCl(aq)$	+	$NaOH(aq)$	→	$NaCl(aq)$	+	$H_2O(l)$

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.

$$\text{moles} = \frac{\text{volume}}{1000} \times \text{concentration}$$
$$= 21.0 \times \frac{0.10}{1000}$$
$$= 2.1 \times 10^{-3}$$

number of moles of hydrochloric acid = number of moles of sodium hydroxide

Therefore, the number of moles of sodium hydroxide
$= 2.1 \times 10^{-3}$

2.1×10^{-3} moles of sodium hydroxide is present in $25.0\,cm^3$ of solution.

Therefore, in $1\,cm^3$ 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 $0.084\,mol\,dm^{-3}$.

You can simplify the calculation by substituting in the following mathematical equation:

$$\frac{M_1 V_1}{M_{acid}} = \frac{M_2 V_2}{M_{alkali}}$$

where:

M_1 = concentration of the acid used
V_1 = volume of acid used (cm^3)
M_{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 (cm^3)
M_{alkali} = number of moles of alkali shown in the chemical equation.

In the example:

$M_1 = 0.10\,mol\,dm^{-3}$
$V_1 = 21.0\,cm^3$
M_{acid} = 1 mole
M_2 = unknown
$V_2 = 25.0\,cm^3$
M_{alkali} = 1 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 $0.084\,mol\,dm^{-3}$.

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, $25\,cm^3$ of it were just neutralised by $20.15\,cm^3$ of a $0.2\,mol\,dm^{-3}$ 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.

sulfuric acid + sodium hydroxide → sodium sulfate + water

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

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 4.03×10^{-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 cm^3 of the solution, so the concentration of the sulfuric acid

$$= 2.015 \times 10^{-3} \times \frac{1000}{25} = 0.081\ mol\,dm^{-3}$$

Questions

1. 24.2 cm^3 of a solution containing 0.20 $mol\,dm^{-3}$ of hydrochloric acid just neutralised 25.0 cm^3 of a potassium hydroxide solution. What is the concentration of this potassium hydroxide solution?
2. 22.4 cm^3 of a solution containing 0.10 $mol\,dm^{-3}$ of sulfuric acid just neutralised 25.0 cm^3 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 $H^+(aq)$ ions as the only positive ions. An acid is a hydrogen ion (H^+) 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_4^+).
- **Alkali** A soluble base which produces $OH^-(aq)$ 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^+) accepter.
- **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 $H^+(aq)$ by reaction with a base, carbonate or metal. Destroying alkalinity means removing the $OH^-(aq)$ by reaction with an acid.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

- **Normal salt** A substance formed when all the replaceable hydrogen of an acid is completely replaced by metal ions or the ammonium ion (NH_4^+).
- **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 $H^+(aq)$ ions in water solution, for example hydrochloric acid.
- **Strong alkali** An acid which produces a high concentration of $OH^-(aq)$ 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 ($CuSO_4.5H_2O$).
- **Weak acid** An acid which produces a low concentration of $H^+(aq)$ in water solution, for example ethanoic acid. It is only partially ionised.
- **Weak alkali** An acid which produces a low concentration of $OH^-(aq)$ in water solution, for example ammonia solution. It is only partially ionised.

Acids, bases and salts

Additional questions

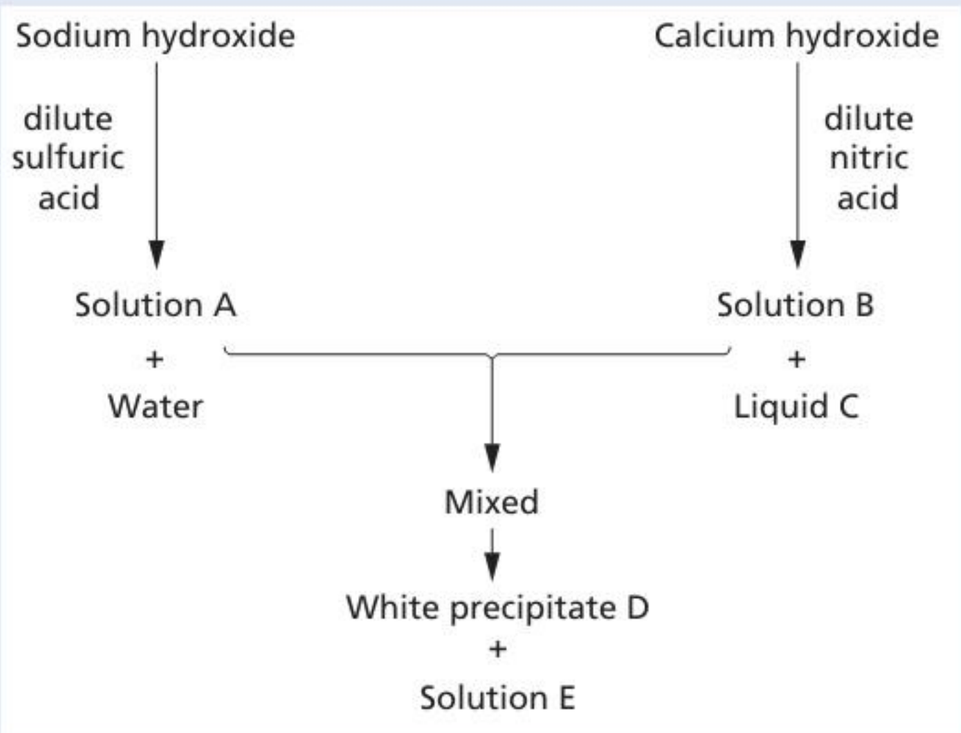
1 Explain, with the aid of examples, what you understand by the following terms:
   a strong acid
   b weak acid
   c strong alkali
   d weak alkali
   e concentrated acid.

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 cm^3 potassium hydroxide solution against a solution containing 1 mol dm^{-3} of sulfuric acid, 28.0 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 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*, $CuSO_4.5H_2O$. 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 $MgSO_4$	= 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 $MgSO_4.xH_2O$, 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.

Period	Group							
	1	2	3	4	5	6	7	8
1	H							
2	Li	Be	B	C	N	O	F	
3	Na	Mg	Al	Si	P	S	Cl	
4	K Cu	Ca Zn	* *	Ti †	V As	Cr Se	Mn Br	Fe Co Ni

Figure 9.2 Mendeleev's periodic table. He left gaps for undiscovered elements.

In 1886 the element we now know as germanium was discovered in Germany by Clemens Winkler; its properties were almost exactly those Mendeleev had predicted. In all, Mendeleev predicted the atomic weight of ten new elements, of which seven were eventually discovered – the other three, atomic weights 45, 146 and 175, do not exist!

The success of Mendeleev's predictions showed that his ideas were probably correct. His periodic table was quickly accepted by scientists as an important summary of the properties of the elements.

Mendeleev's periodic table has been modified in the light of work carried out by Rutherford and Moseley. Discoveries about sub-atomic particles led them to realise that the elements should be arranged by proton number. In the modern Periodic Table the 118 known elements are arranged in order of increasing proton number (Figure 9.3). Those elements with similar chemical properties are found in the same columns or **groups**. There are eight groups of elements. The first column is called Group I; the second Group II; and so on up to Group VII. The final column in the Periodic Table is called Group 0 (or Group VIII). Some of the groups have been given names.

Group I: The alkali metals
Group II: The alkaline earth metals
Group VII: The halogens
Group 0: Inert gases or noble gases

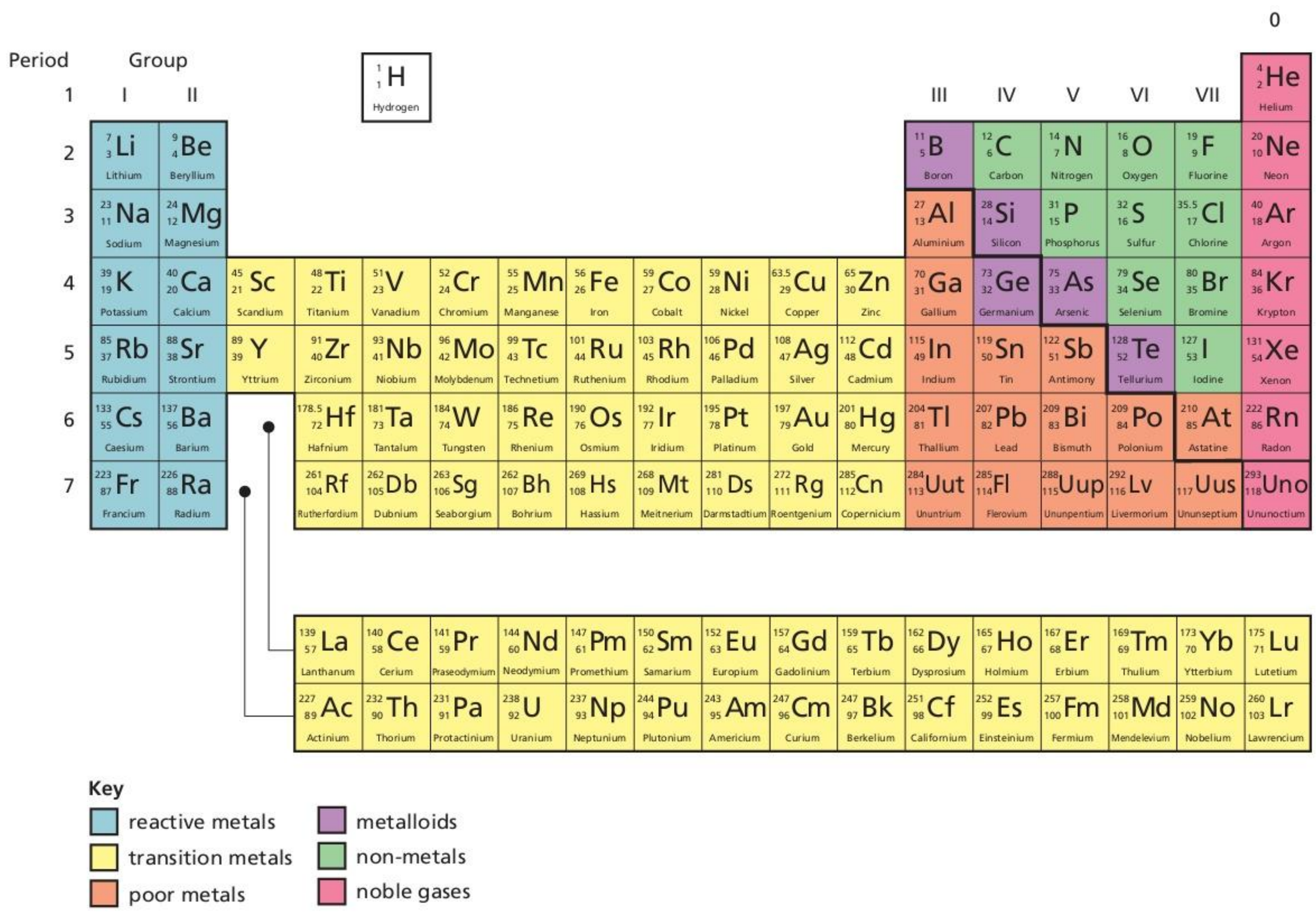


Figure 9.3 The modern Periodic Table.

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).

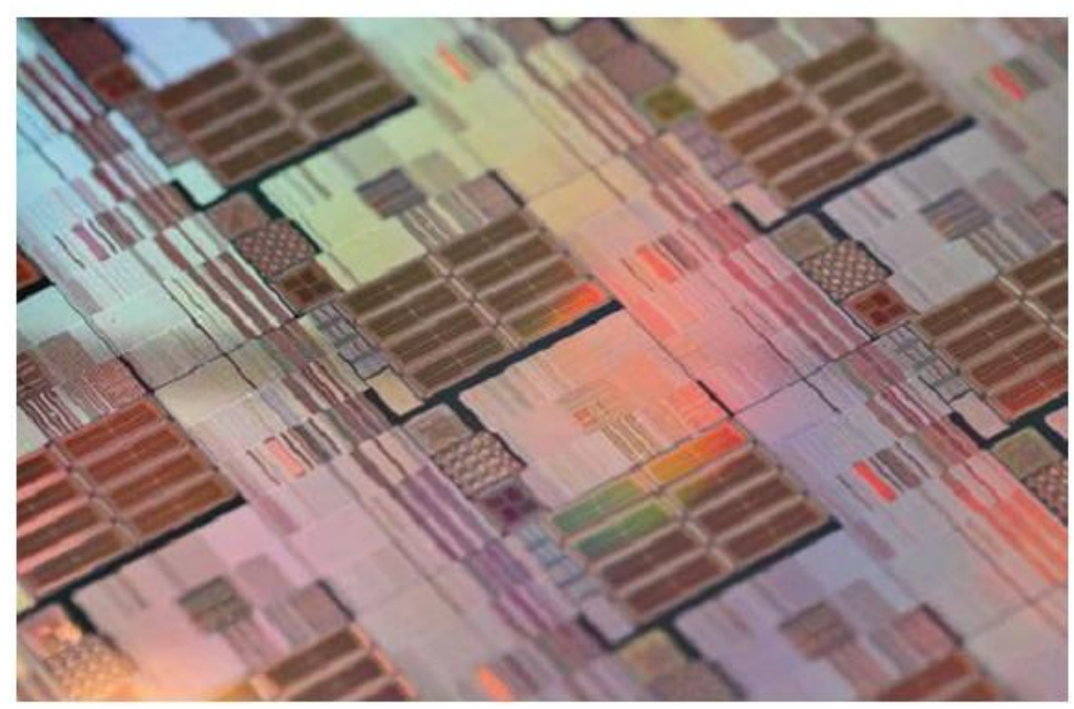


Figure 9.5 The metalloid silicon is used to make silicon 'chips'.

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 $0.53\,g\,cm^{-3}$ and potassium has a density of $0.86\,g\,cm^{-3}$.
- 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 °C and potassium has a melting point of 64 °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:

$$\text{lithium} + \text{oxygen} \rightarrow \text{lithium oxide}$$
$$4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$$

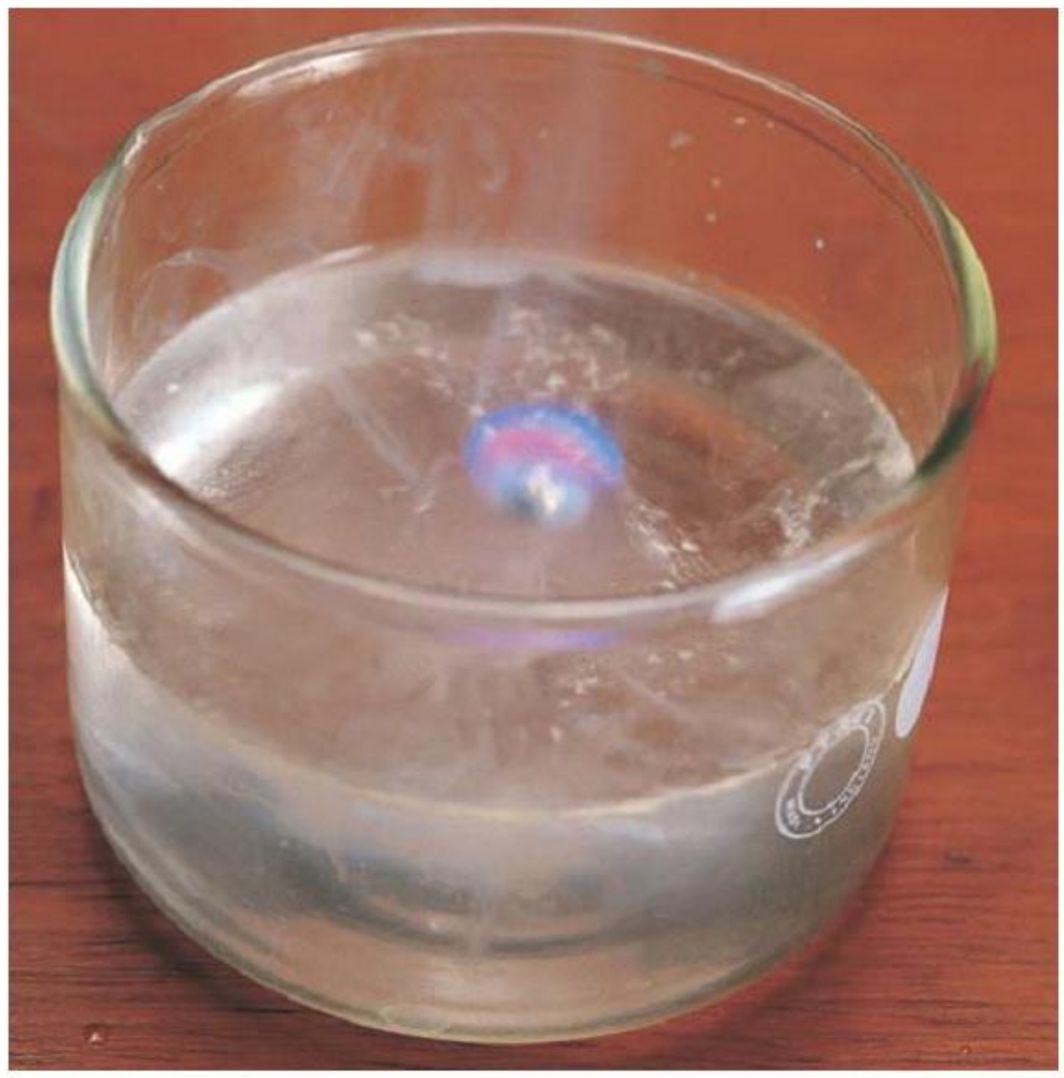
These Group I oxides all dissolve in water to form alkaline solutions of the metal hydroxide.

$$\text{lithium oxide} + \text{water} \rightarrow \text{lithium hydroxide}$$
$$Li_2O(s) + H_2O(l) \rightarrow 2LiOH(aq)$$

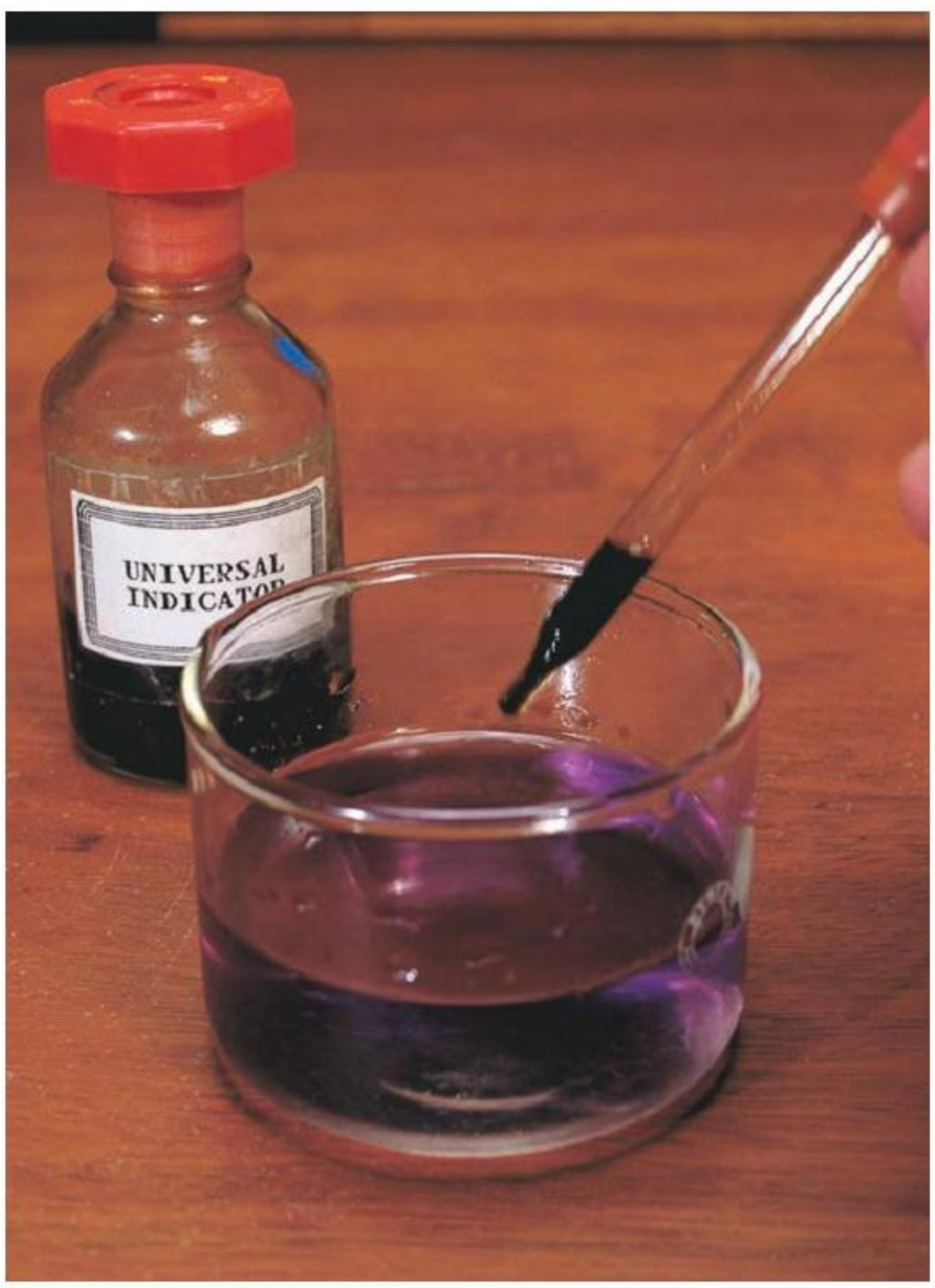
- They react vigorously with water to give an alkaline solution of the metal hydroxide as well as producing hydrogen gas. For example:

$$\text{potassium} + \text{water} \rightarrow \text{potassium hydroxide} + \text{hydrogen gas}$$
$$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$$

- 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).

sodium + chlorine → sodium chloride

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

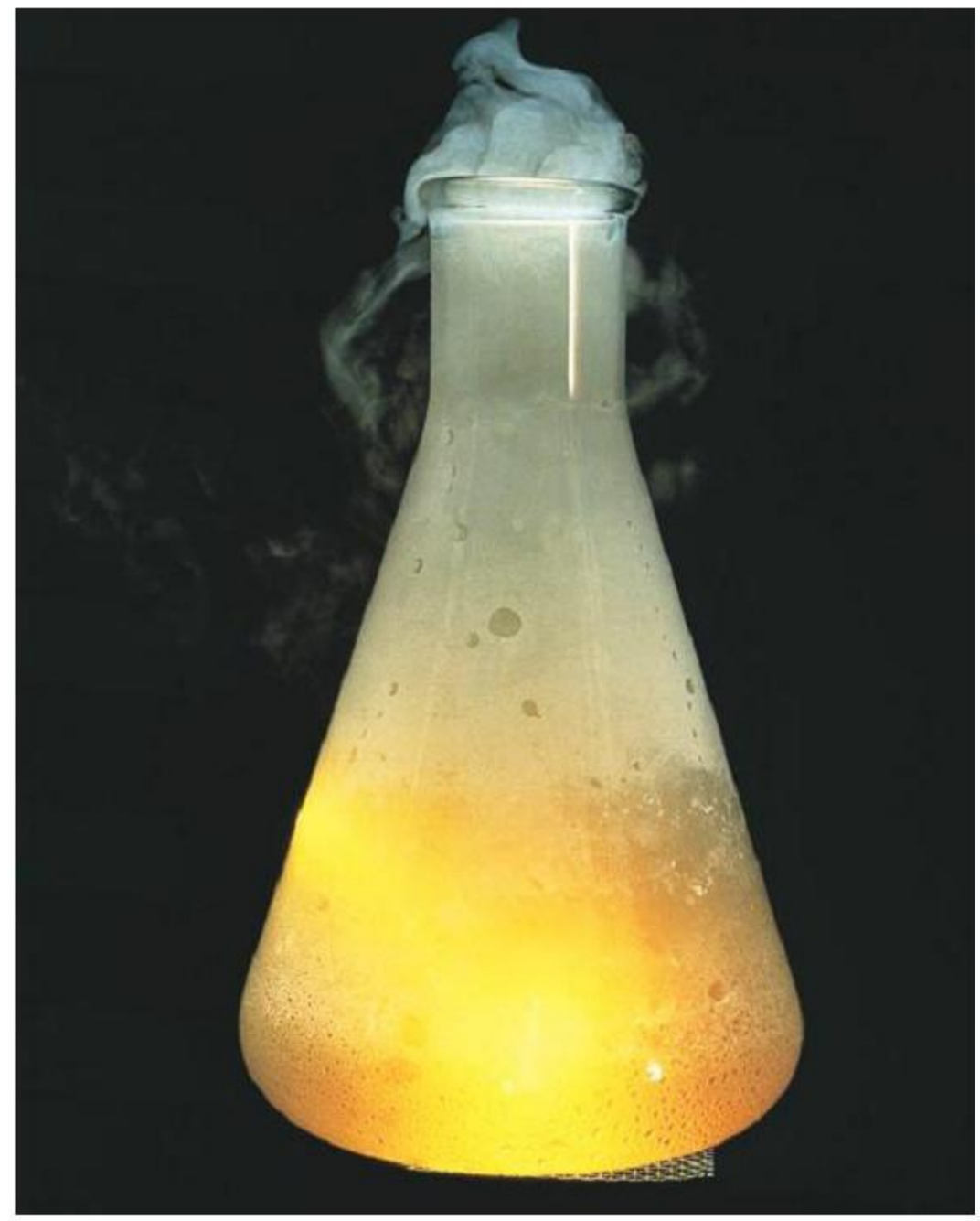


Figure 9.8 A very vigorous reaction takes place when sodium burns in chlorine gas. Sodium chloride is produced.

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).

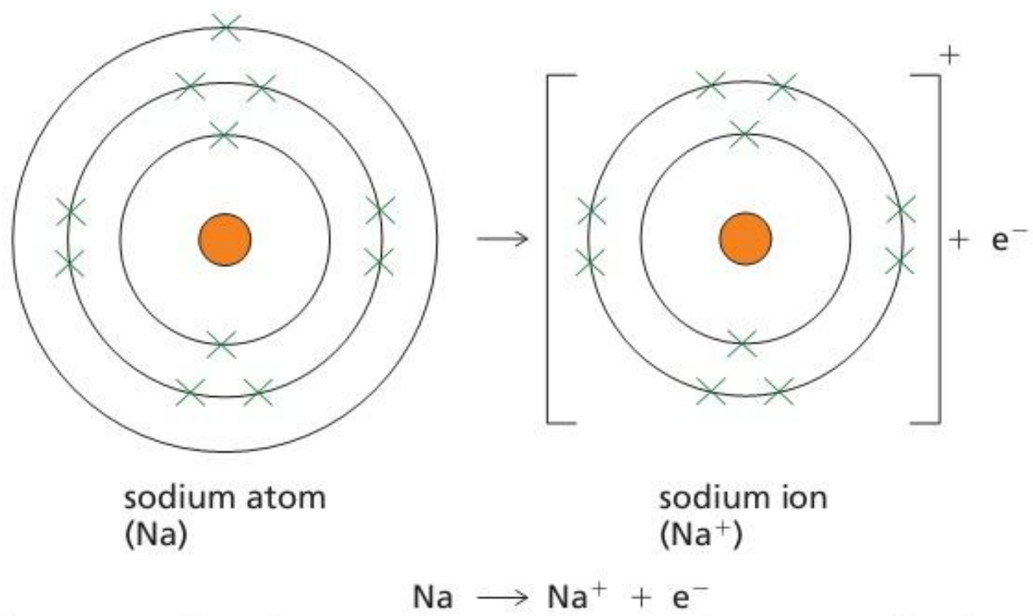


Figure 9.9 This sodium atom loses an electron to become a sodium ion.

Look at Figure 9.10. Why do you think potassium is more reactive than lithium or sodium?

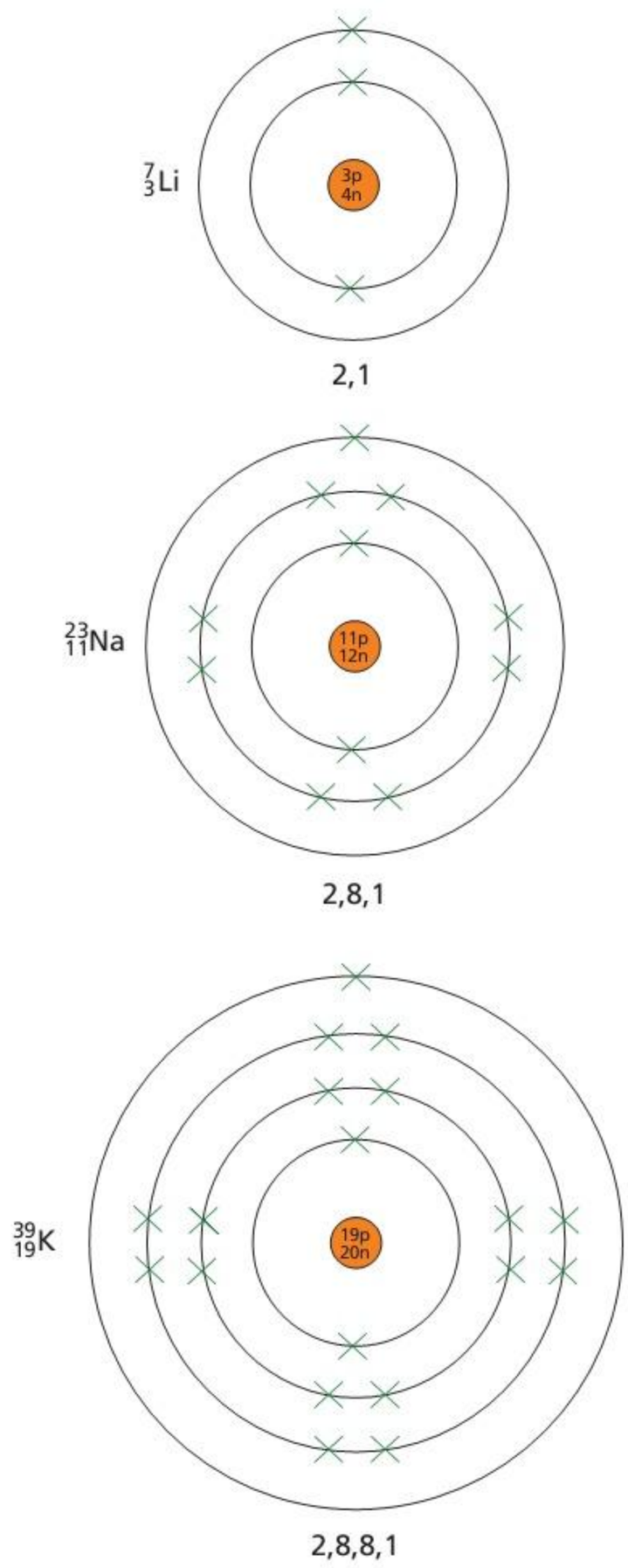


Figure 10 Electronic structures of lithium, sodium and potassium.

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

1 Write word and balanced chemical equations for the reactions between:
   a sodium and oxygen
   b sodium and water.
2 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.
3 Write word and balanced chemical equations for the reactions between:
   a magnesium and water
   b calcium and oxygen.
4 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.

- They are good conductors of heat and electricity.
- They burn in oxygen or air with characteristic flame colours to form solid white oxides. For example:

magnesium + oxygen → magnesium oxide

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

- They react with water, but they do so much less vigorously than the elements in Group I. For example:

calcium + water → calcium hydroxide + hydrogen gas

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$$

Considering the group as a whole, the further down the group you go, the more reactive the elements become.

Group VII – the halogens

Group VII consists of the four elements fluorine, chlorine, bromine and iodine, and the radioactive element astatine. Of these five elements, chlorine, bromine and iodine are generally available for use in school.

- These elements are coloured and darken going down the group (Table 9.4).

Table 9.4 Colours of some halogens.

Halogen	Colour
Chlorine	Pale green
Bromine	Red–brown
Iodine	Purple–black

- They exist as diatomic molecules, for example Cl_2, Br_2 and I_2.
- They show a gradual change from a gas (Cl_2), through a liquid (Br_2), to a solid (I_2) (Figure 9.12) as the density increases.
- They form molecular compounds with other non-metallic elements, for example HCl.
- They react with hydrogen to produce the hydrogen halides, which dissolve in water to form acidic solutions.

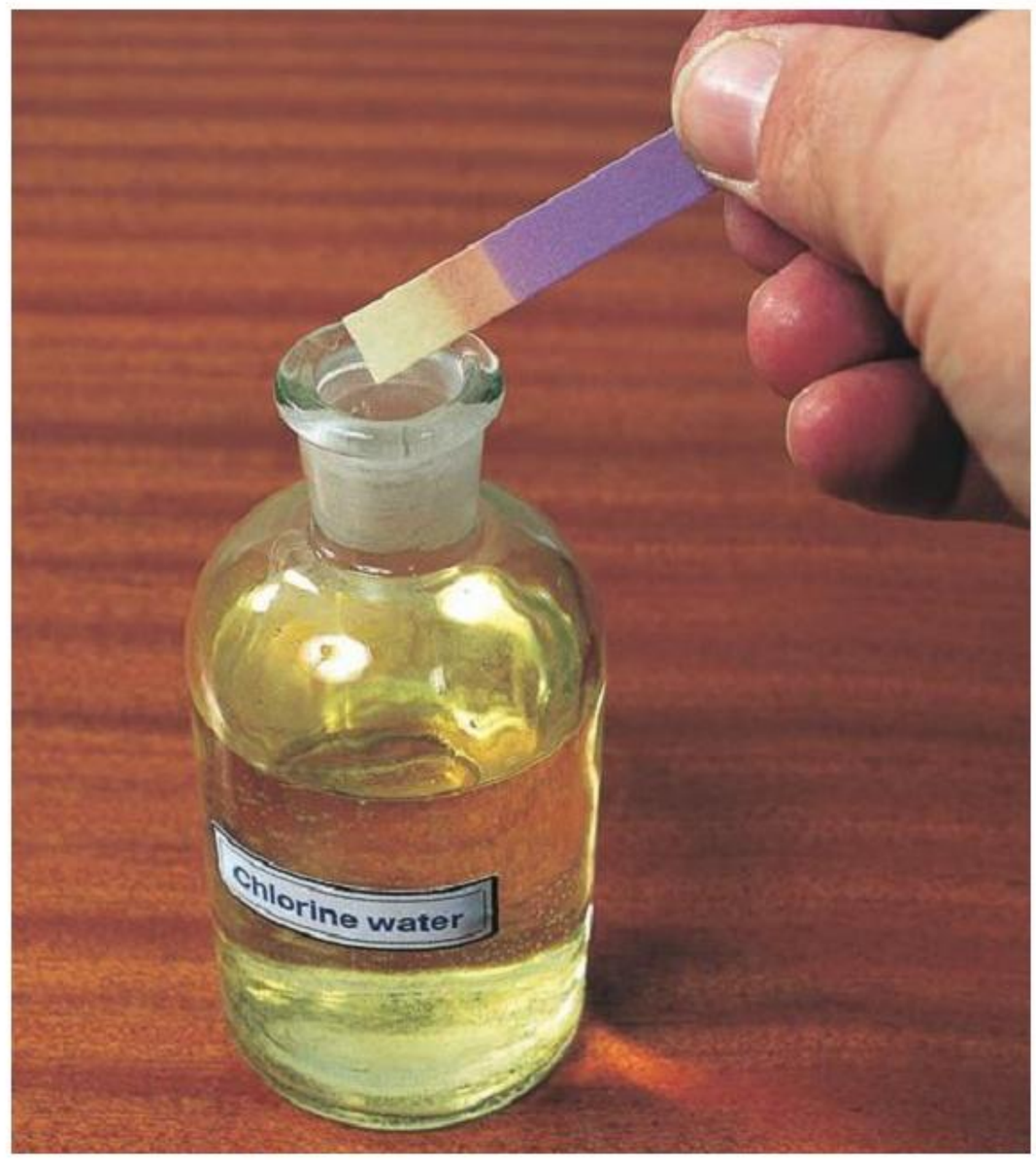
hydrogen + chlorine → hydrogen chloride

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

hydrogen chloride + water → hydrochloric acid

$$HCl(g) + H_2O \rightarrow HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

a Chlorine, bromine and iodine.



b Chlorine gas bleaches moist indicator paper.

Figure 9.12

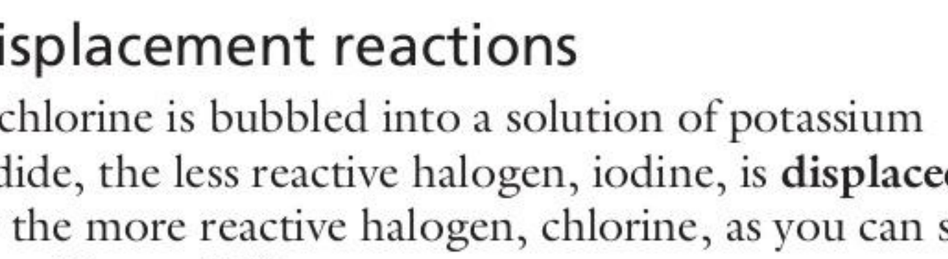
- They react with metals to produce ionic metal halides, for example chlorine and iron produce iron(III) chloride.

iron + chlorine → iron(III) chloride

$$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$$

Displacement reactions

If chlorine is bubbled into a solution of potassium iodide, the less reactive halogen, iodine, is **displaced** by the more reactive halogen, chlorine, as you can see from Figure 9.13:



$$\text{potassium iodide} + \text{chlorine} \rightarrow \text{potassium chloride} + \text{iodine}$$

$$2KI(aq) + Cl_2(g) \rightarrow 2KCl(aq) + I_2(aq)$$

Figure 9.13 Iodine being displaced from potassium iodide solution as chlorine is bubbled through.

The observed order of reactivity of the halogens, confirmed by similar displacement reactions, is:

Decreasing reactivity →

chlorine bromine iodine

You will notice that, in contrast to the elements of Groups I and II, the order of reactivity decreases on going down the group.

Table 9.5 shows the electronic structure for chlorine and bromine. In each case the outer energy level contains seven electrons. When these elements react they gain one electron per atom to gain the stable electron configuration of a noble gas. You will learn more about the stable nature of these gases in the next section. For example, when chlorine reacts it gains a single electron and forms a negative ion (Figure 9.14).

Table 9.5 Electronic structure of chlorine and bromine.

Element	Symbol	Proton number	Electronic structure
Chlorine	Cl	17	2,8,7
Bromine	Br	35	2,8,18,7

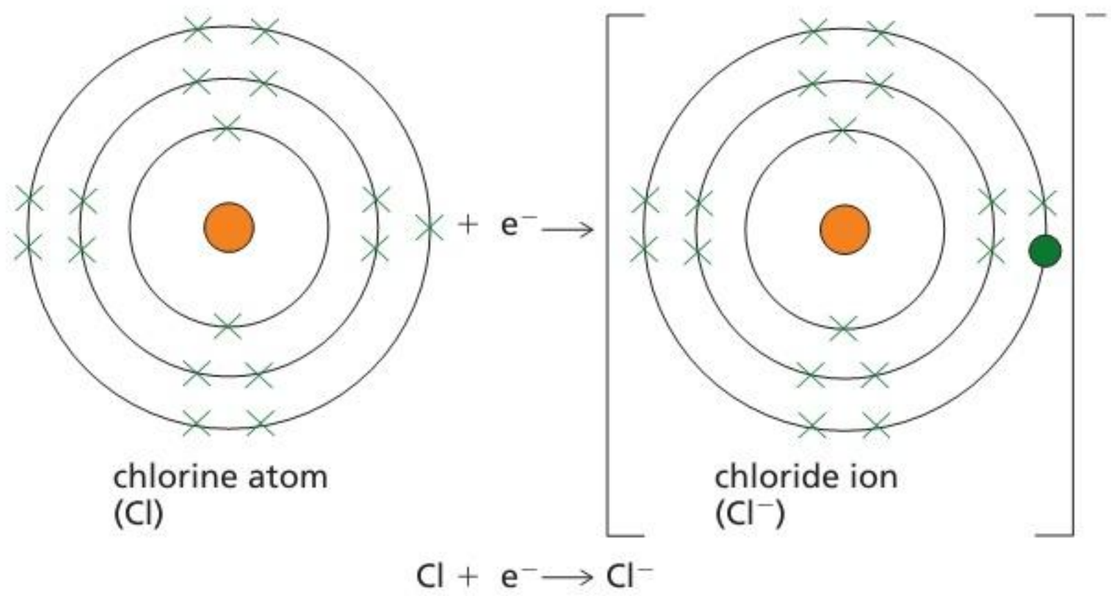


Figure 9.14 A chlorine atom gains an electron to form a chloride ion.

Chlorine is more reactive than bromine because the incoming electron is being more strongly attracted into the outer energy level of the smaller atom. The attractive force on it will be greater than in the case of bromine, since the outer energy level of chlorine is closer to the nucleus. As you go down the group this outermost extra electron is further from the nucleus. It will, therefore, be held less securely, and the resulting reactivity of the elements in Group VII will decrease down the group.

The halogens and halogenic compounds are used in a multitude of different ways (Figure 9.15).

In the reaction of chlorine with potassium iodide, both Cl atoms in Cl_2 gain an electron from an iodide ion, I^-, thus forming two chloride ions, Cl^-. The iodine atoms so formed, by the loss of an electron, combine to give an iodine molecule, I_2.

Oxidation number	$2I^-(aq)$	$+ Cl_2(g) \rightarrow$	$2Cl^-(aq)$	$+ I_2(aq)$
	−1	0	−1	0
Colour	colourless	green	colourless	orange-brown

The iodide ion has been oxidised because it has lost electrons. The oxidation number has increased. Chlorine has been reduced because it has gained electrons. The oxidation number has decreased.

- Fluorine is used in the form of fluorides in drinking water and toothpaste because it reduces tooth decay by hardening the enamel on teeth.
- Chlorine is used to make PVC plastic as well as household bleaches. It is also used to kill bacteria and viruses in drinking water (Chapter 11, p. 191).
- Bromine is used to make disinfectants, medicines and fire retardants.
- Iodine is used in medicines and disinfectants and also as a photographic chemical.

Group 0 – the noble gases

Helium, neon, argon, krypton, xenon and the radioactive element radon make up a most unusual group of non-metals, called the noble gases. They were all discovered after Mendeleev had published his periodic table. They were discovered between 1894 and 1900, mainly through the work of the British scientists Sir William Ramsay (Figure 9.16a) and Lord John William Strutt Rayleigh (Figure 9.16b).

- They are colourless gases.
- They are monatomic gases – they exist as individual atoms, for example He, Ne and Ar.
- They are very unreactive.
- An alternative name for Group 0 is Group VIII

No compounds of helium, neon or argon have ever been found. However, more recently a number of compounds of xenon and krypton with fluorine and oxygen have been produced, for example XeF_6.

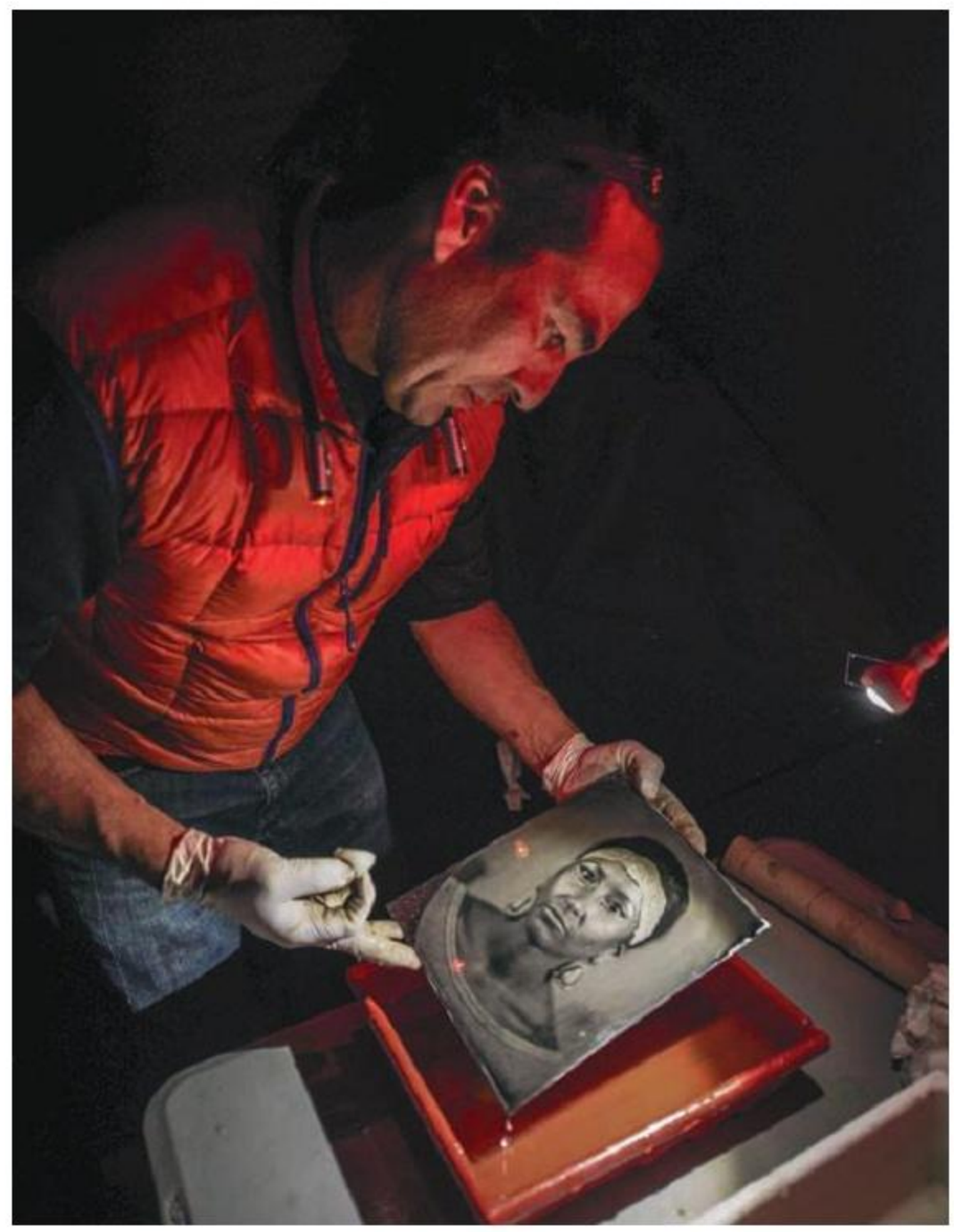
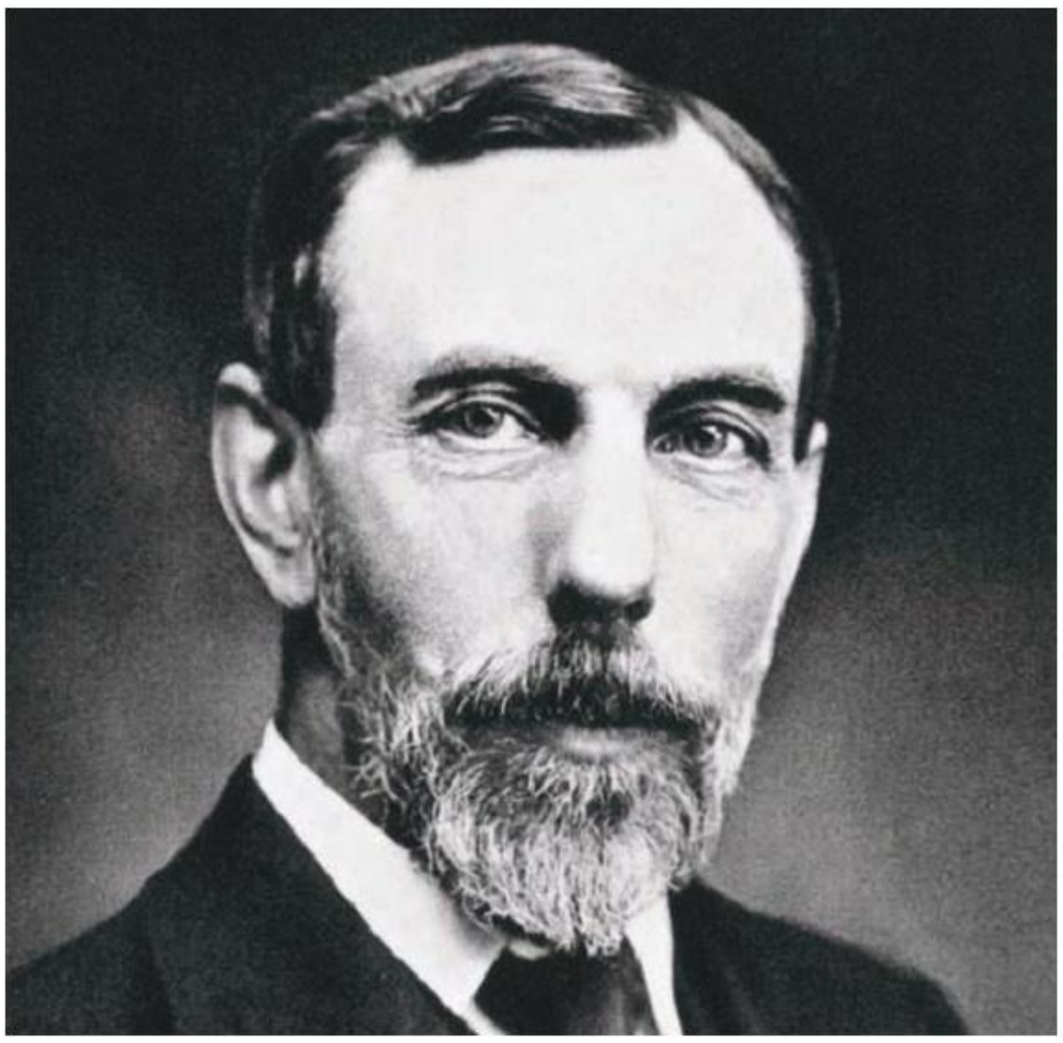


Figure 9.15 The halogens have many varied uses – fluoride in toothpaste to help reduce dental decay, chlorine in household bleach to kill bacteria, bromine as a fire retardant, and iodine in photographic reproduction.



a Sir William Ramsay (1852–1916).

b Lord Rayleigh (1842–1919).

Figure 9.16 Both helped to discover the noble gases and won the Nobel Prize in Chemistry in 1904 for their work.

These gases are chemically unreactive because they have electronic structures which are stable and very difficult to change (Table 9.6). They are so stable that other elements attempt to attain these electron configurations during chemical reactions (Chapter 3, p. 38, and p. 142). You have probably seen this in your study of the elements of Groups I, II and VII.

Table 9.6 Electronic structure of helium, neon and argon.

Element	Symbol	Proton number	Electronic structure
Helium	He	2	2
Neon	Ne	10	2,8
Argon	Ar	18	2,8,8

Although unreactive, they have many uses. Argon, for example, is the gas used to fill light bulbs to prevent the tungsten filament reacting with air. Neon is used extensively in advertising signs and in lasers. Further uses of these gases are discussed in Chapter 11, p. 176.

Helium is separated from natural gas by the liquefaction of the other gases. The other noble gases are obtained in large quantities by the fractional distillation of liquid air (Chapter 11, p. 174).

Questions

1 Write word and balanced chemical equations for the reactions between:
   a bromine and potassium iodide solution
   b bromine and potassium chloride solution.

   If no reaction will take place, write 'no reaction' and explain why.
   c Using the information on pp. 141–142, predict the properties of the element astatine related to its melting point, density and physical state at room temperature. Predict how astatine would react with sodium bromide solution.
2 Write down the names and symbols for the noble gases not given in Table 9.6 and use your research skills to find a use for each.

Transition elements

This block of metals includes many you will be familiar with, for example copper, iron, nickel, zinc and chromium (Figure 9.17).

- They are harder and stronger than the metals in Groups I and II.
- They have much higher densities than the metals in Groups I and II.
- They have high melting points (except for mercury, which is a liquid at room temperature).
- They are less reactive metals.
- They form a range of brightly coloured compounds (Figure 9.18, p. 146).
- They are good conductors of heat and electricity.
- They show catalytic activity (Chapter 7, p. 109) as elements and compounds. For example, iron is used in the industrial production of ammonia gas (Haber process, Chapter 11, p. 177).
- They do not react (corrode) so quickly with oxygen and/or water.

a Copper is used in many situations which involve good heat and electrical conduction. It is also used in medallions and bracelets.

b These gates are made of iron. Iron can easily be moulded into different shapes.

c Monel is an alloy of nickel and copper. It is extremely resistant to corrosion, even that caused by sea water.

d This bucket has been coated with zinc to prevent the steel of the bucket corroding.

e The alloy stainless steel contains a high proportion of chromium, which makes it corrosion resistant.

Figure 9.17 Everyday uses of transition elements and their compounds. They are often known as the 'everyday metals'.

- They form simple ions with more than one oxidation state. (For a discussion of oxidation states see Chapter 3, p. 43.) For example, copper forms Cu^+ (Cu(I)) and Cu^{2+} (Cu(II)), in compounds such as Cu_2O and $CuSO_4$, and iron forms Fe^{2+} (Fe(II)) and Fe^{3+} (Fe(III)), in compounds such as $FeSO_4$ and $FeCl_3$.
- They form more complicated ions with high oxidation states. For example, chromium forms the dichromate(VI) ion, $Cr_2O_7^{2-}$, which contains chromium with a +6 oxidation state (Cr(VI)) and manganese forms the manganate(VII) ion, MnO_4^-, which contains manganese with a +7 oxidation state (Mn(VII)).

a Some solutions of coloured transition element compounds.

b The coloured compounds of transition elements can be seen in these pottery glazes.

Figure 9.18

Questions

1 Look at the photographs in Figure 9.17 (p. 145) and decide which properties are important when considering the particular use the metal is being put to.

2 Which groups in the Periodic Table contain:
   a only metals?
   b only non-metals?
   c both metals and non-metals?

The position of hydrogen

Hydrogen is often placed by itself in the Periodic Table. This is because the properties of hydrogen are unique. However, profitable comparisons can be made with the other elements. It is often shown at the top of either Group I or Group VII, but it cannot fit easily into the trends shown by either group; see Table 9.7.

Table 9.7 Comparison of hydrogen with lithium and fluorine.

Lithium	Hydrogen	Fluorine
Solid	Gas	Gas
Forms a positive ion	Forms positive or negative ions	Forms a negative ion
1 electron in outer energy level	1 electron in outer energy level	1 electron short of a full outer energy level
Loses 1 electron to form a noble gas configuration	Needs 1 electron to form a noble gas configuration	Needs 1 electron to form a noble gas configuration

Checklist

After studying Chapter 9 you should know and understand the following terms.

- **Alkali metals** The six metallic elements found in Group I of the Periodic Table.
- **Alkaline earth metals** The six metallic elements found in Group II of the Periodic Table.
- **Displacement reaction** A reaction in which a more reactive element displaces a less reactive element from solution.
- **Group** A vertical column of the Periodic Table containing elements with similar properties with the same number of electrons in their outer energy levels. They have an increasing number of inner energy levels as you descend the group.
- **Halogens** The elements found in Group VII of the Periodic Table.
- **Metalloid (semi-metal)** Any of the class of chemical elements intermediate in properties between metals and non-metals, for example boron and silicon.
- **Noble gases** The unreactive gases found in Group 0 of the Periodic Table.
- **Periodic Table** A table of elements arranged in order of increasing proton number to show the similarities of the chemical elements with related electronic structures.
- **Periods** Horizontal rows of the Periodic Table. Within a period the atoms of all the elements have the same number of occupied energy levels but have an increasing number of electrons in the outer energy level.
- **Transition elements** The elements found in the centre of the Periodic Table, between Groups II and III.

The Periodic Table

Additional questions

1 The diagram below shows part of the Periodic Table.

I	II											III	IV	V	VI	VII	0
			H														He
Li	Be											B			O		
													Si	P		Cl	Ar
	Ca										Zn	Ga			Se	Br	

Using only the symbols of the elements shown above, write down the symbol for an element which:

a is a pale green coloured toxic gas
b is stored under oil
c has five electrons in its outer electron energy shell
d is the most reactive Group II element
e is the most reactive halogen
f is the only liquid shown
g is a transition element
h is a gas with two electrons in its outer shell.

2 The proton number of barium (Ba) is 56. It is in Group II of the Periodic Table.

a How many electrons would you expect a barium atom to contain in its outer energy level?
b How would you expect barium to react with chlorine? Write a word and balanced chemical equation for this reaction.
c How would you expect barium to react with water? Write a word and balanced chemical equation for this reaction.
d Write down the formulae of the bromide and sulfate of barium.

3 Find the element germanium (Ge) in the Periodic Table.

a Which group of the Periodic Table is this element in?
b How many electrons will it have in its outer energy level (shell)?
c Is germanium a metal or a non-metal?
d What is the formula of the chloride of germanium?
e Name and give the symbols of the other elements in this group.

4 Three members of the halogens are: $^{35.5}_{17}Cl$, $^{80}_{35}Br$ and $^{127}_{53}I$.

a **(i)** Write down the electronic structure of an atom of chlorine.
(ii) Why is the relative atomic mass of chlorine not a whole number?
(iii) How many protons are there in an atom of bromine?
(iv) How many neutrons are there in an atom of iodine?
(v) State and account for the order of reactivity of these elements.

b When potassium is allowed to burn in a gas jar of chlorine, in a fume cupboard, clouds of white smoke are produced.
(i) Why is this reaction carried out in a fume cupboard?
(ii) What does the white smoke consist of?
(iii) Write a word and balanced chemical equation for this reaction.
(iv) Describe what you would expect to see when potassium is allowed to burn safely in a gas jar of bromine vapour. Write a word and balanced chemical equation for this reaction.

5 'By using displacement reactions it is possible to deduce the order of reactivity of the halogens.' Discuss this statement with reference to the elements bromine, iodine and chlorine only.

6 Use the information given in the table below to answer the questions below concerning the elements **Q**, **R**, **S**, **T** and **X**.

Element	Proton number	Nucleon number	Electronic structure
Q	3	7	2,1
R	20	40	2,8,8,2
S	18	40	2,8,8
T	8	18	2,6
X	19	39	2,8,8,1

a Which element has 22 neutrons in each atom?
b Which element is a noble gas?
c Which two elements form ions with the same electronic structure as neon?
d Which two elements are in the same group of the Periodic Table and which group is this?
e Place the elements in the table into the periods in which they belong.
f Which is the most reactive metal element shown in the table?
g Which of the elements is calcium?

7 **a** Consider the chemical properties and physical properties of the halogens chlorine, bromine and iodine. Using these properties, predict the following about the other two halogens, fluorine and astatine.

Property	Fluorine	Astatine
State at room temperature and pressure		
Colour		
Reactivity with sodium metal		

b (i) Write a word equation for the reaction of fluorine gas with sodium chloride solution.
(ii) Write a balanced chemical equation for the reaction, with state symbols.
(iii) Write an ionic equation for the reaction, with state symbols.

8 Some of the most important metals we use are found in the transition element section of the Periodic Table. One of these elements is copper. Sodium, a Group I metal, has very different properties from those of copper. Complete the table below to show their differences.

	Transition element, e.g. copper	Group I metal, e.g. sodium
Hardness (hard/soft)		
Reactivity		
Density (high/low)		
Catalytic activity		
Variable oxidation states		

10 Metals

Metal reactions
With acid
With air/oxygen
With water/steam

Decomposition of metal nitrates, carbonates, oxides and hydroxides
Metal nitrates
Metal carbonates
Metal oxides
Metal hydroxides

Reactivity of metals and their uses
Using the reactivity series
Competition reactions in the solid state
Competition reactions in aqueous solutions

Identifying metal ions
Amphoteric hydroxides and oxides

Discovery of metals and their extraction
Extraction of reactive metals
Extraction of fairly reactive metals
Extraction of unreactive metals

Metal waste

Rusting of iron
Rust prevention
Corrosion

Alloys
Production of steel
Alloys to order

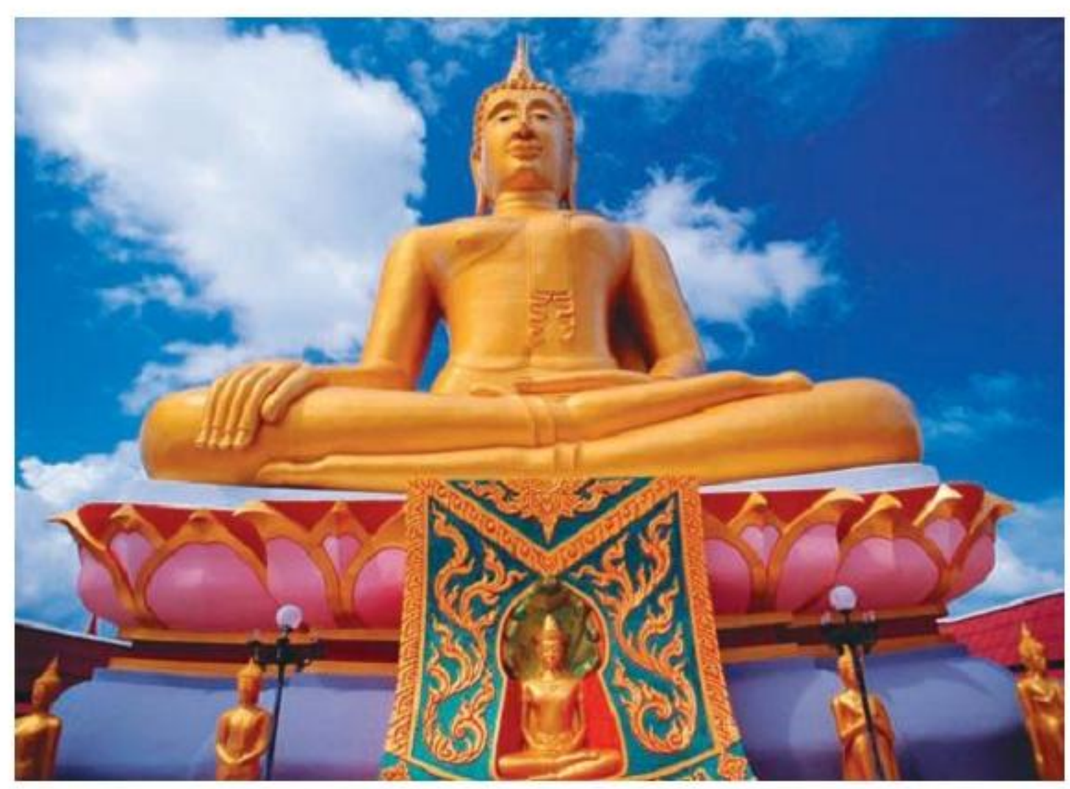
Checklist

Additional questions

You have already seen in Chapter 2, p. 12, that metals usually have similar physical properties. However, they differ in other ways. Look closely at the three photographs in Figure 10.1.

a Sodium burning in air/oxygen.

b Iron rusts when left unprotected.



c Gold is used in leaf form on this giant Buddha as it is unreactive.

Figure 10.1

Sodium is soft and reacts violently with both air and water. Iron also reacts with air and water but much more slowly, forming rust. Gold, however, remains totally unchanged after many hundreds of years. Sodium is said to be more reactive than iron and, in turn, iron is said to be more reactive than gold.

Metal reactions

By carrying out reactions in the laboratory with other metals and with air, water and dilute acid, it is possible to produce an order of reactivity of the metals.

With acid

Look closely at the photograph in Figure 10.2 showing magnesium metal reacting with dilute hydrochloric acid. You will notice effervescence, which is caused by bubbles of hydrogen gas being formed as the reaction between the two substances proceeds. The other product of this reaction is the salt, magnesium chloride.

magnesium + hydrochloric acid → magnesium chloride + hydrogen

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

If a metal reacts with dilute hydrochloric acid then hydrogen and the metal chloride are produced.

If similar reactions are carried out using other metals with acid, an order of reactivity can be produced by measuring the rate of evolution of hydrogen. This is known as a **reactivity series**.

Figure 10.2 Effervescence occurs when magnesium is put into acid.

An order of reactivity, giving the most reactive metal first, using results from experiments with dilute acid, is shown in Table 10.1. The table also shows how the metals react with air/oxygen and water/steam, and, in addition, the ease of extraction of the metal.

In all these reactions the most reactive metal is the one that has the highest tendency to lose outer electrons to form a positive metal ion.

Table 10.1 Order of reactivity.

Reactivity series	Reaction with dilute acid	Reaction with air/oxygen	Reaction with water	Ease of extraction
Potassium (K)	Produce H_2 with decreasing vigour	Burn very brightly and vigorously	Produce H_2 with decreasing vigour with cold water	Difficult to extract
Sodium (Na)				
Calcium (Ca)		Burn to form an oxide with decreasing vigour	React with steam with decreasing vigour	Easier to extract
Magnesium (Mg)				
Aluminium (Al*)				
[Carbon (C)]				
Zinc (Zn)				
Iron (Fe)				
Lead (Pb)		React slowly to form the oxide	Do not react with cold water or steam	
[Hydrogen (H)]				
Copper (Cu)	Do not react with dilute acids			
Silver (Ag)		Do not react		Found as the element (native)
Gold (Au)				
Platinum (Pt)				

Increasing reactivity of metal

* Because aluminium reacts so readily with the oxygen in the air, a protective oxide layer is formed on its surface. This often prevents any further reaction and disguises aluminium's true reactivity. This gives us the use of a light and strong metal.

With air/oxygen

Many metals react directly with oxygen to form oxides. For example, magnesium burns brightly in oxygen to form the white powder magnesium oxide.

magnesium + oxygen → magnesium oxide

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

With water/steam

Reactive metals such as potassium, sodium and calcium react with cold water to produce the metal hydroxide and hydrogen gas. For example, the reaction of sodium with water produces sodium hydroxide and hydrogen.

sodium + water → sodium hydroxide + hydrogen

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

The moderately reactive metals, magnesium, zinc and iron, react slowly with water. They will, however, react more rapidly with steam (Figure 10.3). In their reaction with steam, the metal oxide and hydrogen are formed. For example, magnesium produces magnesium oxide and hydrogen gas.

magnesium + steam → magnesium oxide + hydrogen

$$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$$

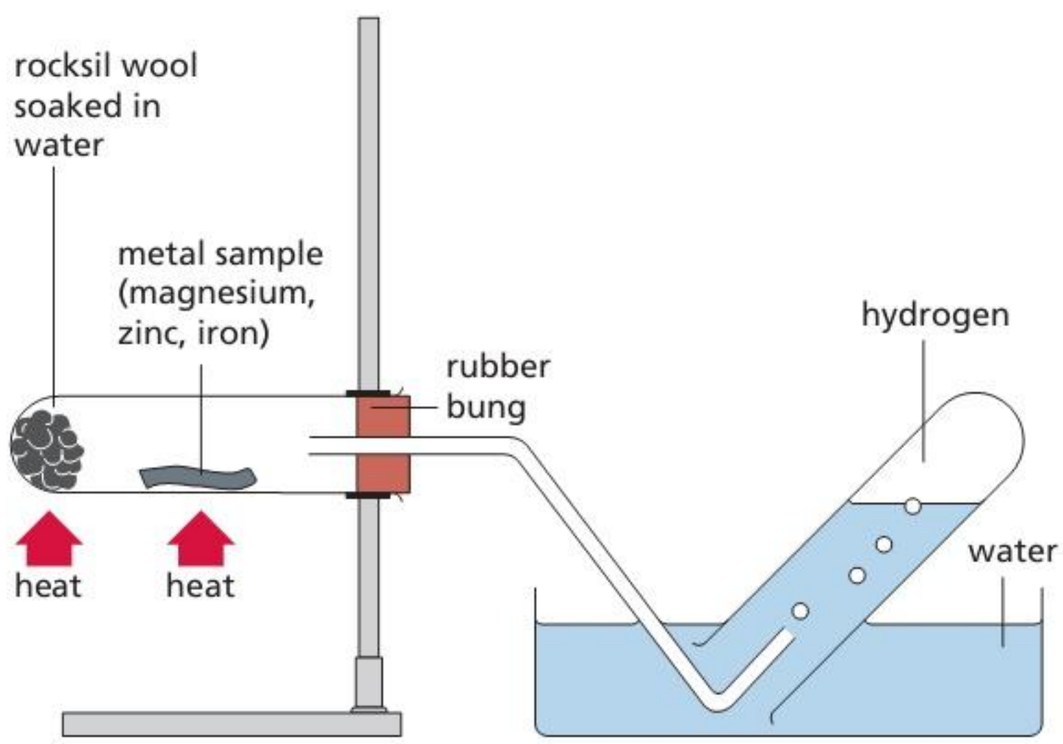


Figure 10.3 Apparatus used to investigate how metals such as magnesium react with steam.

Generally, it is the unreactive metals that we find the most uses for; for example, the metals iron and copper can be found in many everyday objects (Figure 10.4). However, magnesium is one of the metals used in the construction of the Airbus A380 (Figure 10.5).

a This wood-burning stove is made of iron.

b Copper pots and pans.

Figure 10.4

Figure 10.5 Planes are made of an alloy which contains magnesium and aluminium.

Both sodium and potassium are so reactive that they have to be stored under oil to prevent them from coming into contact with water or air. However, because they have low melting points and are good conductors of heat, they are used as coolants for nuclear reactors.

Questions

1 Write balanced chemical equations for the reactions between:
   a iron and dilute hydrochloric acid
   b zinc and oxygen
   c calcium and water.
2 Make a list of six things you have in your house made from copper or iron.
   Give a use for each of the other unreactive metals shown in the reactivity series.

Decomposition of metal nitrates, carbonates, oxides and hydroxides

The ways in which metal nitrates, carbonates, oxides and hydroxides decompose can also be discussed in terms of the reactivity series of the metals. The decomposition processes are different, depending on the position of the metal in the reactivity series.

Metal nitrates

When nitrates of reactive metals are heated, they decompose to produce the metal nitrite and oxygen gas.

$$\text{sodium nitrate} \xrightarrow{\text{heat}} \text{sodium nitrite} + \text{oxygen}$$
$$2NaNO_3(s) \longrightarrow 2NaNO_2(s) + O_2(g)$$

Nitrates of moderately reactive metals produce brown fumes of nitrogen dioxide gas when heated, as well as the metal oxide and oxygen gas.

$$\text{magnesium nitrate} \xrightarrow{\text{heat}} \text{magnesium oxide} + \text{nitrogen dioxide} + \text{oxygen}$$
$$2Mg(NO_3)_2(s) \longrightarrow 2MgO(s) + 4NO_2(g) + O_2(g)$$

Different products again are produced when nitrates of unreactive metals are heated. These metal nitrates decompose to give the metal, nitrogen dioxide gas and oxygen.

Metal carbonates

All metal carbonates that exist undergo thermal decomposition to give the metal oxide and carbon dioxide gas. The general equation is:

$$\text{metal carbonate} \xrightarrow{\text{heat}} \text{metal oxide} + \text{carbon dioxide}$$

Carbonates of reactive metals such as potassium and sodium do not easily undergo any decomposition reaction. The carbonates of these metals are thermally stable and require very high temperatures to force them to decompose.

Carbonates of moderately reactive metals do decompose. The less reactive the metal, the lower the temperature needed to make the carbonate decompose.

The carbonates of unreactive metals such as gold, silver and platinum are too unstable to exist.

Metal oxides

Theoretically, all metal oxides can be thermally decomposed to give the metal and oxygen gas. In reality, it is usually too difficult to decompose the oxides of reactive metals, and even some moderately reactive metals, such as aluminium, have oxides that require thousands of degrees to make them decompose.

It is possible to thermally decompose some oxides of less reactive metals such as silver oxide.

$$\text{silver oxide} \xrightarrow{\text{heat}} \text{silver} + \text{oxygen gas}$$
$$2Ag_2O(s) \longrightarrow 4Ag(s) + O_2(g)$$

Table 10.2 Effect of heat on metal nitrates.

Metal	Symbol	Effect of heat on nitrate	Equation (some examples)
Potassium Sodium	K Na	Form the metal nitrite and oxygen is released	potassium nitrate → potassium nitrite + oxygen $2KNO_3(s) \rightarrow 2KNO_2(s) + O_2(g)$
Calcium Magnesium Aluminium Zinc Iron Lead Copper	Ca Mg Al Zn Fe Pb Cu	Form the oxide and release oxygen Brown fumes of nitrogen dioxide (nitrogen(IV) oxide) are also produced	calcium nitrate → calcium oxide + nitrogen(IV) oxide + oxygen $2Ca(NO_3)_2(s) \rightarrow 2CaO(s) + 4NO_2(g) + O_2(g)$ zinc nitrate → zinc oxide + nitrogen(IV) oxide + oxygen $2Zn(NO_3)_2(s) \rightarrow 2ZnO(s) + 4NO_2(g) + O_2(g)$ copper(II) nitrate → copper oxide + nitrogen(IV) oxide + oxygen $2Cu(NO_3)_2(s) \rightarrow 2CuO(s) + 4NO_2(g) + O_2(g)$
Silver Platinum Gold	Ag Pt Au	Decompose to the metal releasing oxygen and nitrogen(IV) oxide	silver nitrate → silver + nitrogen(IV) oxide + oxygen $2AgNO_3(s) \rightarrow 2Ag(s) + 2NO_2(g) + O_2(g)$

Table 10.3 Effect of heat on metal carbonates.

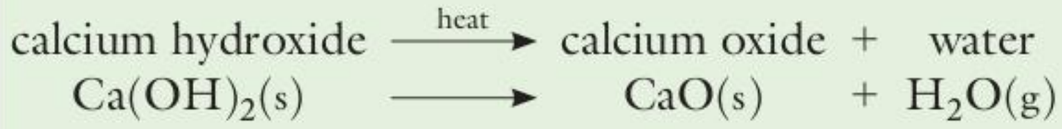
Metal	Symbol	Effect of heat on carbonate	Equation (some examples)
Potassium Sodium	K Na	No reaction	
Calcium Magnesium Aluminium Zinc Lead Copper	Ca Mg Al Zn Pb Cu	Form the oxide and release carbon dioxide with increasing ease	calcium carbonate → calcium oxide + carbon dioxide $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ zinc carbonate → zinc oxide + carbon dioxide $ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$ copper(II) carbonate → copper(II) oxide + carbon dioxide $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$
Silver Platinum Gold	Ag Pt Au	Carbonates are too unstable to exist	

Table 10.4 Effect of heat on metal hydroxides.

Metal	Symbol	Effect of heat on hydroxide	Equation (some examples)
Potassium Sodium	K Na	No reaction (i.e. there is no decomposition)	
Calcium Magnesium Aluminium Zinc Iron Lead Copper	Ca Mg Al Zn Fe Pb Cu	Decompose to the oxide and release water	calcium hydroxide → calcium oxide + water $Ca(OH)_2(s) \rightarrow CaO(s) + H_2O(g)$ zinc hydroxide → zinc oxide + water $Zn(OH)_2(s) \rightarrow ZnO(s) + H_2O(g)$ copper hydroxide → copper oxide + water $Cu(OH)_2(s) \rightarrow CuO(s) + H_2O(g)$
Silver Platinum Gold	Ag Pt Au	Hydroxides are too unstable to exist	

Metal hydroxides

Hydroxides of reactive metals show no decomposition when they are heated. The hydroxides of moderately reactive metals do decompose to produce the metal oxide and water. This process is used to convert calcium hydroxide (slaked lime) into calcium oxide (lime).



Reactivity of metals and their uses

Generally, it is the unreactive metals for which we find most uses. For example, the metals iron and copper can be found in everyday objects such as car bodies and coins, respectively. However, the metal aluminium is an exception. Aluminium appears in the reactivity series just below magnesium and is quite reactive. Fortunately, it forms a relatively thick oxide layer on the surface of the metal which prevents further reaction. This gives us the use of a light, strong metal for items such as door handles and cooking foil. There are many reactive metals you would not expect to find many uses for. For example, sodium is so reactive it is stored under oil to prevent it coming into contact with water and the air. However, because it has a low melting point and is a good conductor of heat, it can be used as a coolant in nuclear reactors.

Question

1 Choose a reactive metal (e.g. magnesium) and an unreactive metal (e.g. lead) and compare the uses of these metals with their reactivity and properties.

Using the reactivity series

The reactivity series is useful for predicting how metals will react. It can also be used to predict the reactions of some metal compounds. The tables on

pp. 152–153 show how some metal nitrates and hydroxides behave when heated. You can see that metal compounds from a similar part of the series behave in a similar manner.

Competition reactions in the solid state

A more reactive metal has a greater tendency to form a metal ion by losing electrons than a less reactive metal does. Therefore, if a more reactive metal is heated with the oxide of a less reactive metal, then it will remove the oxygen from it (as the oxide anion). You can see from the reactivity series that iron is less reactive than aluminium (p. 150). If iron(III) oxide is mixed with aluminium and the mixture is heated using a magnesium fuse (Figure 10.6), a very violent reaction occurs as the competition between the aluminium and the iron for the oxygen takes place.

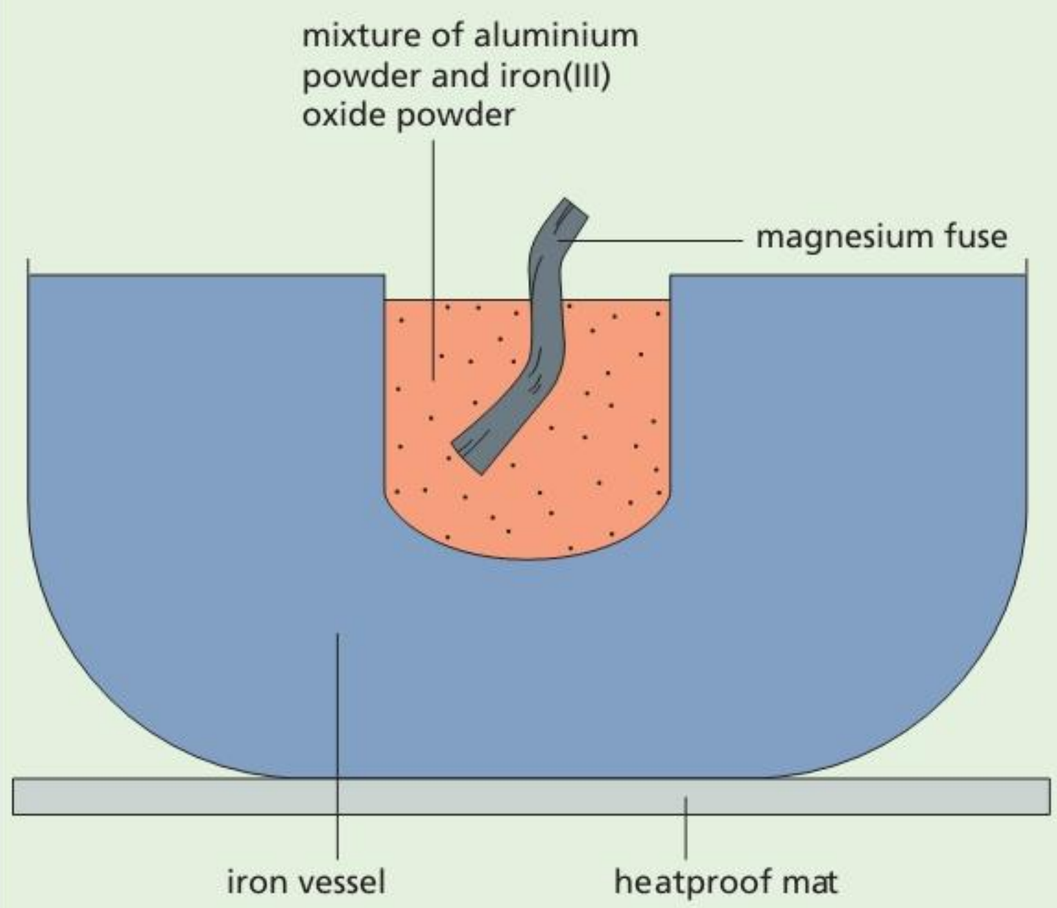


Figure 10.6 Thermit reaction mixture apparatus.

The aluminium, being the more reactive metal, takes the oxygen from the less reactive iron. It is a very exothermic reaction. When the reaction is over, a solid lump of iron is left along with a lot of white aluminium oxide powder.

$$\text{iron(III) oxide} + \text{aluminium} \xrightarrow{\text{heat}} \text{aluminium oxide} + \text{iron}$$

$$Fe_2O_3(s) + 2Al(s) \longrightarrow Al_2O_3(s) + 2Fe(s)$$

or

$$2Fe^{3+}(s) + 2Al(s) \rightarrow 2Al^{3+}(s) + 2Fe(s)$$

This is a **redox reaction** (see Chapter 2, p. 14 and Chapter 5, p. 73 for a discussion of this type of reaction).

This particular reaction is known as the **Thermit reaction** (Figure 10.7). Since large amounts of heat are given out and the iron is formed in a molten state, this reaction is used to weld together damaged railway lines. It is also used in incendiary bombs.

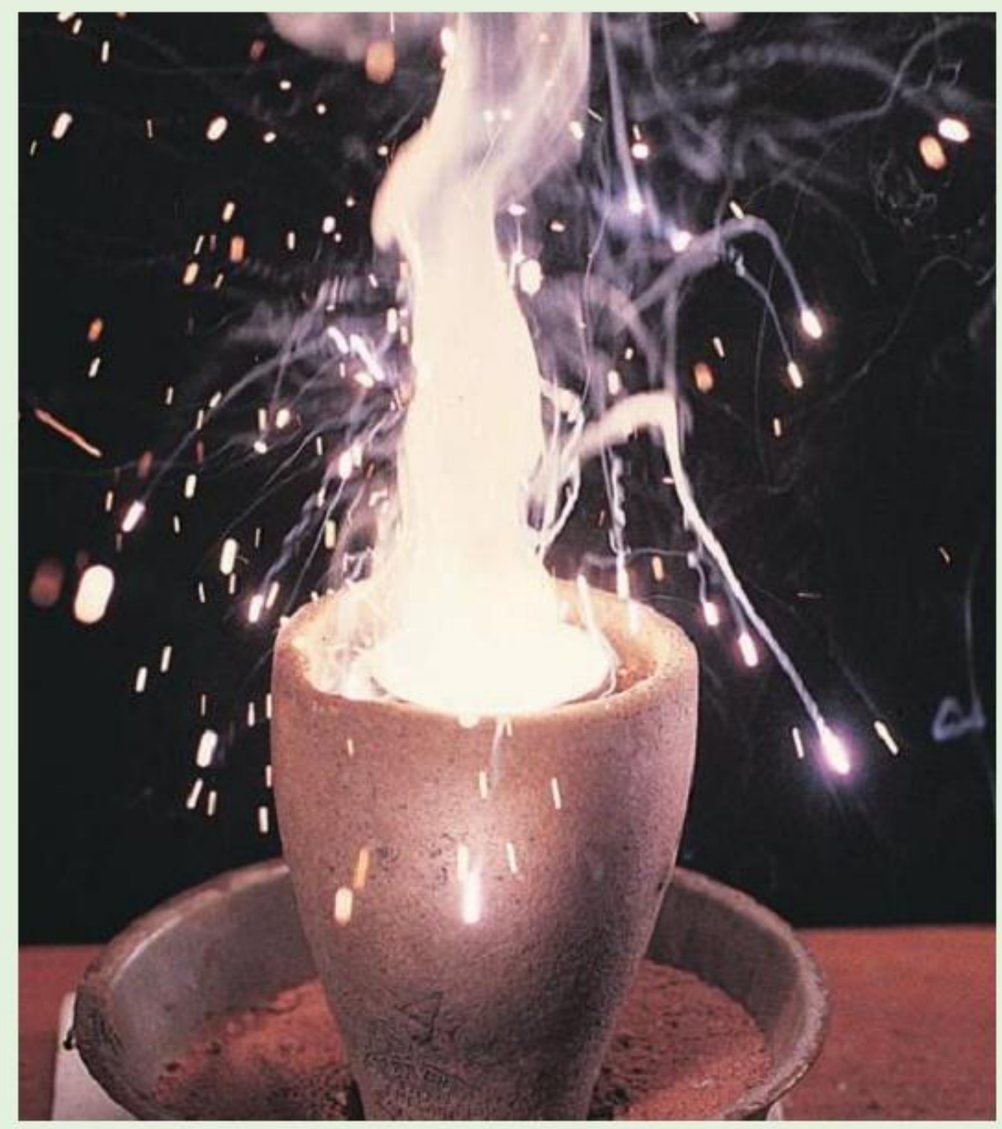


Figure 10.7 The Thermit reaction in a laboratory. The same reaction is used to weld damaged railway lines.

Some metals, such as the transition metals chromium and titanium, are prepared from their oxides using this type of competition reaction.

Carbon, a non-metal, is included in Table 10.1 (p. 150) just below aluminium. It is able to reduce metal oxides below it in the series.

Competition reactions in aqueous solutions

In another reaction, metals compete with each other for other anions. This type of reaction is known as a **displacement reaction**. As in the previous type of competitive reaction, the reactivity series can be used to predict which of the metals will 'win'.

In a displacement reaction, a more reactive metal will displace a less reactive metal from a solution of its salt. Zinc is above copper in the reactivity series.

Figure 10.8 shows what happens when a piece of zinc metal is left to stand in a solution of copper(II) nitrate. The copper(II) nitrate slowly loses its blue colour as the zinc continues to displace the copper from the solution and eventually becomes colourless zinc nitrate.

zinc + copper(II) nitrate → zinc nitrate + copper

$$Zn(s) + Cu(NO_3)_2(aq) \rightarrow Zn(NO_3)_2(aq) + Cu(s)$$

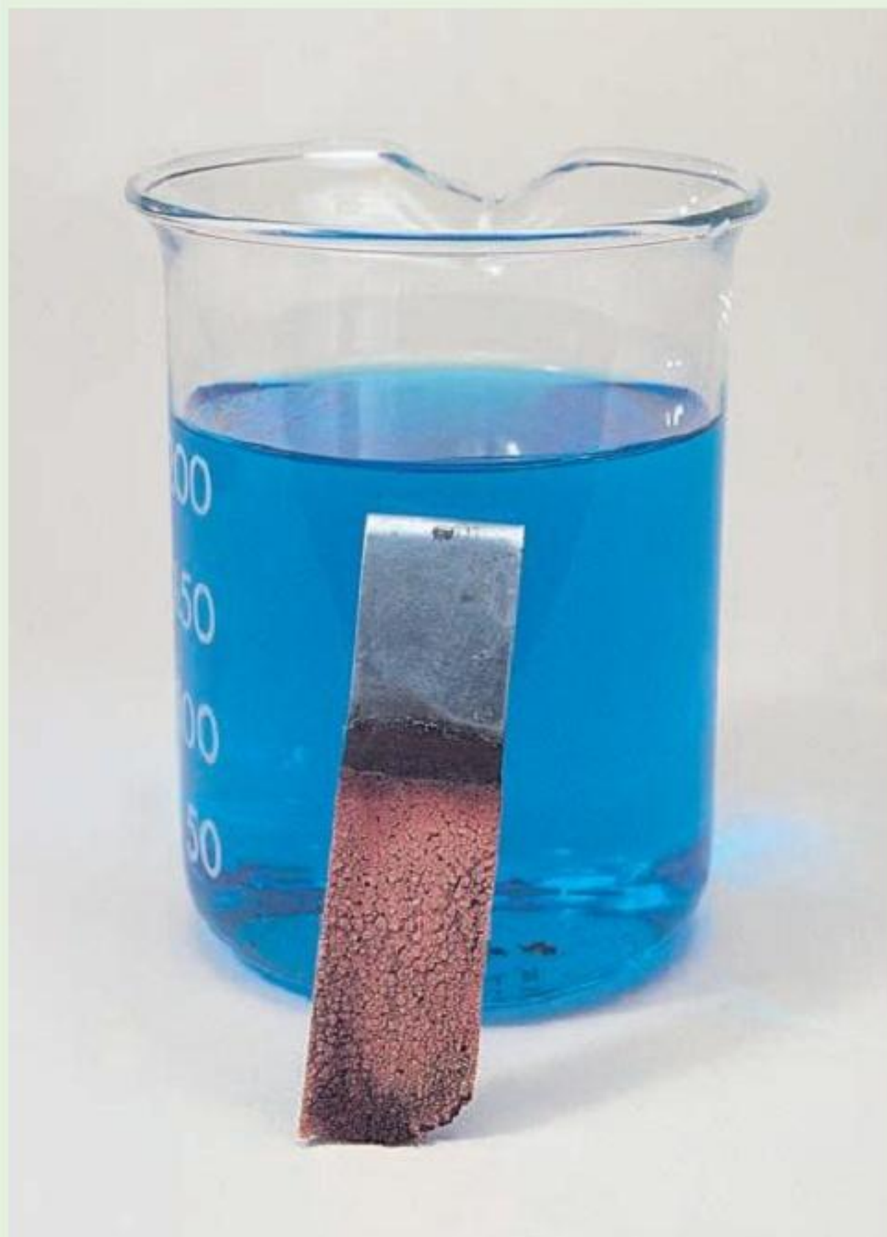


Figure 10.8 Zinc displaces copper.

The ionic equation for this reaction is:

zinc + copper ions → zinc ions + copper

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

This is also a redox reaction involving the transfer of two electrons from the zinc metal to the copper ions. The zinc is oxidised to zinc ions in aqueous solution, while the copper ions are reduced. (See Chapter 5, p. 73, for a discussion of oxidation and reduction in terms of electron transfer.) It is possible to confirm the reactivity series for metals using competition reactions of the types discussed in this section.

Questions

1 Predict whether or not the following reactions will take place:
   a magnesium + copper(II) oxide
   b iron + aluminium oxide
   c calcium + magnesium oxide.
   Complete the word equations and write balanced chemical and ionic equations for those reactions which do take place.

2 Predict whether or not the following reactions will take place:
   a magnesium + calcium nitrate solution
   b iron + copper(II) nitrate solution
   c copper + silver nitrate solution.
   Complete the word equations and write balanced chemical and ionic equations for those reactions which do take place.

3 How could you carry out a series of reactions between metals and solutions of their nitrates to establish a reactivity series?

4 The data below was obtained by carrying out displacement reactions of five metals with the nitrates of the same five metals. Strips of each metal were placed in solutions of the other four metals' nitrate solutions.

	Nitrates				
	A	B	C	D	E
A	—	✓	✓	✓	✗
B	✗	—	✗	✗	✗
C	✗	✓	—	✓	✗
D	✗	✓	✗	—	✗
E	✓	✓	✓	✓	—

✓ = metal displaced ✗ = no reaction

Put the five metals **A**–**E** in order of their reactivity using the data above.

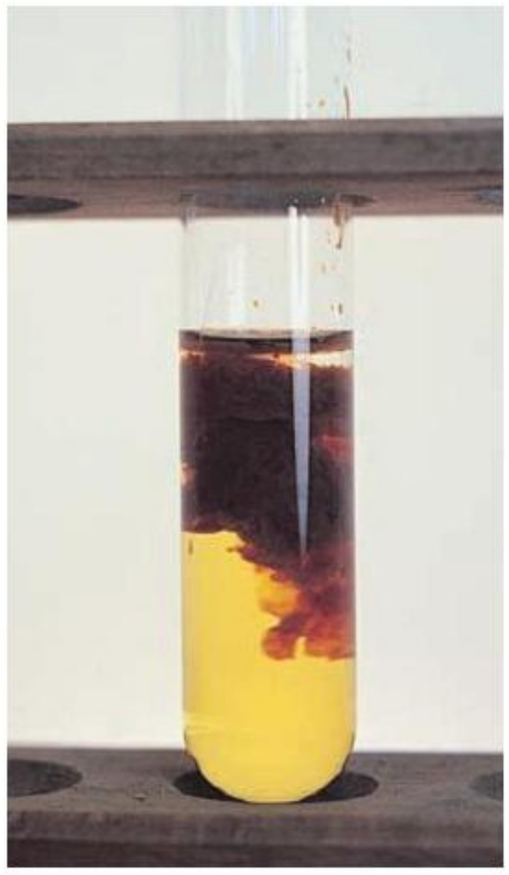
Identifying metal ions

When an alkali dissolves in water, it produces hydroxide ions. It is known that most metal hydroxides are insoluble. So if hydroxide ions from a solution of an alkali are added to a solution of a metal salt, an insoluble, often coloured, metal hydroxide is precipitated from solution (Figure 10.9).

Let's take the example of iron(III) chloride with sodium hydroxide solution:

iron(III) chloride + sodium hydroxide → iron(III) hydroxide + sodium chloride

$$FeCl_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + 3NaCl(aq)$$



a Iron(III) hydroxide is precipitated. **b** Copper(II) hydroxide is precipitated.
Figure 10.9

The ionic equation for this reaction is:

iron(III) ions + hydroxide ions → iron(III) hydroxide

$$Fe^{3+}(aq) + 3OH^{-}(aq) \rightarrow Fe(OH)_3(s)$$

Table 10.5 shows the effects of adding a few drops of sodium hydroxide solution to solutions containing various metal ions, and of adding an excess. The colours of the insoluble metal hydroxides can be used to identify the metal cations present in solution. In some cases the precipitate dissolves in excess hydroxide, owing to the amphoteric nature of the metal hydroxide. This amphoteric nature can also be used to help identify metals such as aluminium and zinc.

Table 10.5 The effect of adding sodium hydroxide solution to solutions containing various metal ions.

Metal ion present in solution	Effect of adding sodium hydroxide solution	
	A few drops	**An excess**
Aluminium	White precipitate of aluminium hydroxide	Precipitate dissolves
Calcium	White precipitate of calcium hydroxide	Precipitate does not dissolve
Copper(II)	Blue precipitate of copper(II) hydroxide	Precipitate does not dissolve
Iron(II)	Green precipitate of iron(II) hydroxide	Precipitate does not dissolve
Iron(III)	Brown precipitate of iron(III) hydroxide	Precipitate does not dissolve
Zinc	White precipitate of zinc hydroxide	Precipitate dissolves
Chromium(III)	Light green precipitate	Precipitate dissolves

Amphoteric hydroxides and oxides

The hydroxides of metals are basic and they react with acids to form salts (Chapter 8, p. 124). The hydroxides of some metals, however, will also react with strong bases, such as sodium hydroxide, to form soluble salts. Hydroxides of this type are said to be **amphoteric**. For example,

zinc hydroxide + hydrochloric acid → zinc chloride + water

$$Zn(OH)_2(aq) + 2HCl(aq) \rightarrow ZnCl_2(aq) + 2H_2O(l)$$

and

zinc hydroxide + sodium hydroxide → sodium zincate

$$Zn(OH)_2(aq) + 2NaOH(aq) \rightarrow Na_2Zn(OH)_4(aq)$$

Other amphoteric hydroxides are lead hydroxide ($Pb(OH)_2$) and aluminium hydroxide ($Al(OH)_3$). We can use this sort of behaviour to help identify metal cations, as their hydroxides are soluble in strong bases.

Both aluminium and zinc metals will also react readily with moderately concentrated acids and alkalis.

zinc + hydrochloric acid → zinc chloride + hydrogen

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

zinc + sodium hydroxide + water → sodium zincate + hydrogen

$$Zn(s) + 2NaOH(aq) + 2H_2O(l) \rightarrow Na_2Zn(OH)_4(aq) + H_2(g)$$

It should be noted that the oxides of the metals used as examples above are also amphoteric. Aluminium oxide and zinc oxide will react with both acids and alkalis.

zinc oxide + hydrochloric acid → zinc chloride + water

$$ZnO(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$$

zinc oxide + sodium hydroxide + water → sodium zincate

$$ZnO(s) + 2NaOH(aq) + H_2O(l) \rightarrow Na_2Zn(OH)_4(aq)$$

Certain non-metal oxides are classified as neutral. These oxides do not react with acids or bases. Examples include carbon monoxide (CO), nitrogen(II) oxide (nitrogen monoxide, NO), and nitrogen(I) oxide (nitrous oxide or dinitrogen oxide, N_2O).

Questions

1 Write equations for the reaction between:
   a aluminium and moderately concentrated hydrochloric acid
   b aluminium and moderately concentrated sodium hydroxide (producing sodium aluminate, $NaAl(OH)_4$).
2 Write ionic equations for the reactions which take place to produce the metal hydroxides shown in Table 10.5.
3 Describe what you would see when sodium hydroxide is added slowly to a solution containing iron(II) nitrate.

Discovery of metals and their extraction

Metals have been used since prehistoric times. Many primitive iron tools have been excavated. These were probably made from small amounts of native iron found in rock from meteorites. It was not until about 2500 BC that iron became more widely used. This date marks the dawn of the iron age, when people learned how to get iron from its ores in larger quantities by reduction using charcoal. An ore is a naturally occurring mineral from which a metal can be extracted.

Over the centuries other metals, which like iron are also relatively low in the reactivity series, were isolated in a similar manner. These included copper, lead, tin and zinc. However, due to the relatively low abundance of the ores containing these metals, they were not extracted and used in large amounts.

Metals high in the reactivity series have proved very difficult to isolate. It was not until more recent times, through Sir Humphry Davy's work on electrolysis, that potassium (1807), sodium (1807), calcium (1808) and magnesium (1808) were isolated. Aluminium, the most plentiful reactive metal in the Earth's crust, was not extracted from its ore until 1827, by Friedrich Wöhler (p. 74), and the extremely reactive metal rubidium was not isolated until 1861 by Robert Bunsen and Gustav Kirchhoff.

The majority of metals are too reactive to exist on their own in the Earth's crust, and they occur naturally in rocks as compounds in ores (Figure 10.10). These ores are usually carbonates, oxides or sulfides of the metal, mixed with impurities.

Some metals, such as gold and silver, occur in a native form as the free metal (Figure 10.11). They are very unreactive and have withstood the action of water and the atmosphere for many thousands of years without reacting to become compounds.

Some of the common ores are shown in Table 10.6.

Figure 10.10 Metal ores – chalcopyrite (top) and galena.

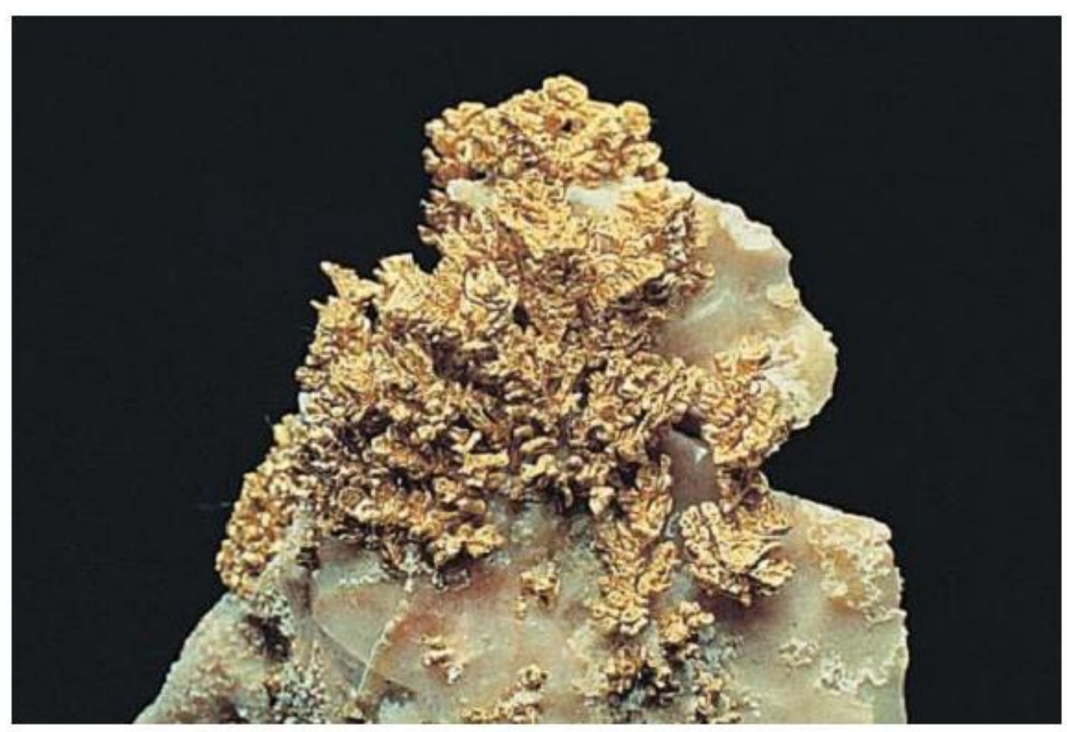


Figure 10.11 Gold crystals.

Large lumps of the ore are first crushed and ground up by very heavy machinery. Some ores are already fairly concentrated when mined. For example, in some parts of the world, haematite contains over 80% Fe_2O_3. However, other ores, such as copper pyrites, are often found to be less concentrated, with only 1% or less of the copper compound, and so they have to be concentrated before the metal can be extracted. The method used to extract the metal from its ore depends on the position of the metal in the reactivity series.

Table 10.6 Some common ores.

Metal	Name of ore	Chemical name of compound in ore	Formula	Usual method of extraction
Aluminium	Bauxite	Aluminium oxide	$Al_2O_3.2H_2O$	Electrolysis of oxide dissolved in molten cryolite
Copper	Copper pyrites	Copper iron sulfide	$CuFeS_2$	The sulfide ore is roasted in air
Iron	Haematite	Iron(III) oxide	Fe_2O_3	Heat oxide with carbon
Sodium	Rock salt	Sodium chloride	NaCl	Electrolysis of molten sodium chloride
Zinc	Zinc blende	Zinc sulfide	ZnS	Sulfide is roasted in air and the oxide produced is heated with carbon

Extraction of reactive metals

Because reactive metals, such as sodium, hold on to the element(s) they have combined with, they are usually difficult to extract. For example, sodium chloride (as rock salt) is an ionic compound with the Na^+ and Cl^- ions strongly bonded to one another. The separation of these ions and the subsequent isolation of the sodium metal is therefore difficult.

Electrolysis of the molten, purified ore is the method used in these cases. During this process, the metal is produced at the cathode while a non-metal is produced at the anode. As you might expect, extraction of metal by electrolysis is expensive. In order to keep costs low, many metal smelters using electrolysis are situated in regions where there is hydroelectric power (Chapter 6, p. 94).

For further discussion of the extraction of aluminium, see Chapter 5, pp. 74–76.

Extraction of fairly reactive metals

Metals towards the middle of the reactivity series, such as iron and zinc, may be extracted by reducing the metal oxide with the non-metal carbon.

Iron

Iron is extracted mainly from its oxides, haematite (Fe_2O_3) and magnetite (Fe_3O_4), in a blast furnace (Figures 10.12 and 10.13). These ores contain at least 60% iron. The iron ores used are a blend of those extracted in Australia, Canada, Sweden, Venezuela and Brazil. The blast furnace is a steel tower approximately 50 m high lined with heat-resistant bricks. It is loaded with the 'charge' of iron ore (usually haematite), coke (made by heating coal) and limestone (calcium carbonate).

A blast of hot air is sent in near the bottom of the furnace through holes (tuyères) which makes the 'charge' glow, as the coke burns in the preheated air.

$$\text{carbon} + \text{oxygen} \rightarrow \text{carbon dioxide}$$
$$C(s) + O_2(g) \rightarrow CO_2(g)$$

Figure 10.12 A blast furnace.

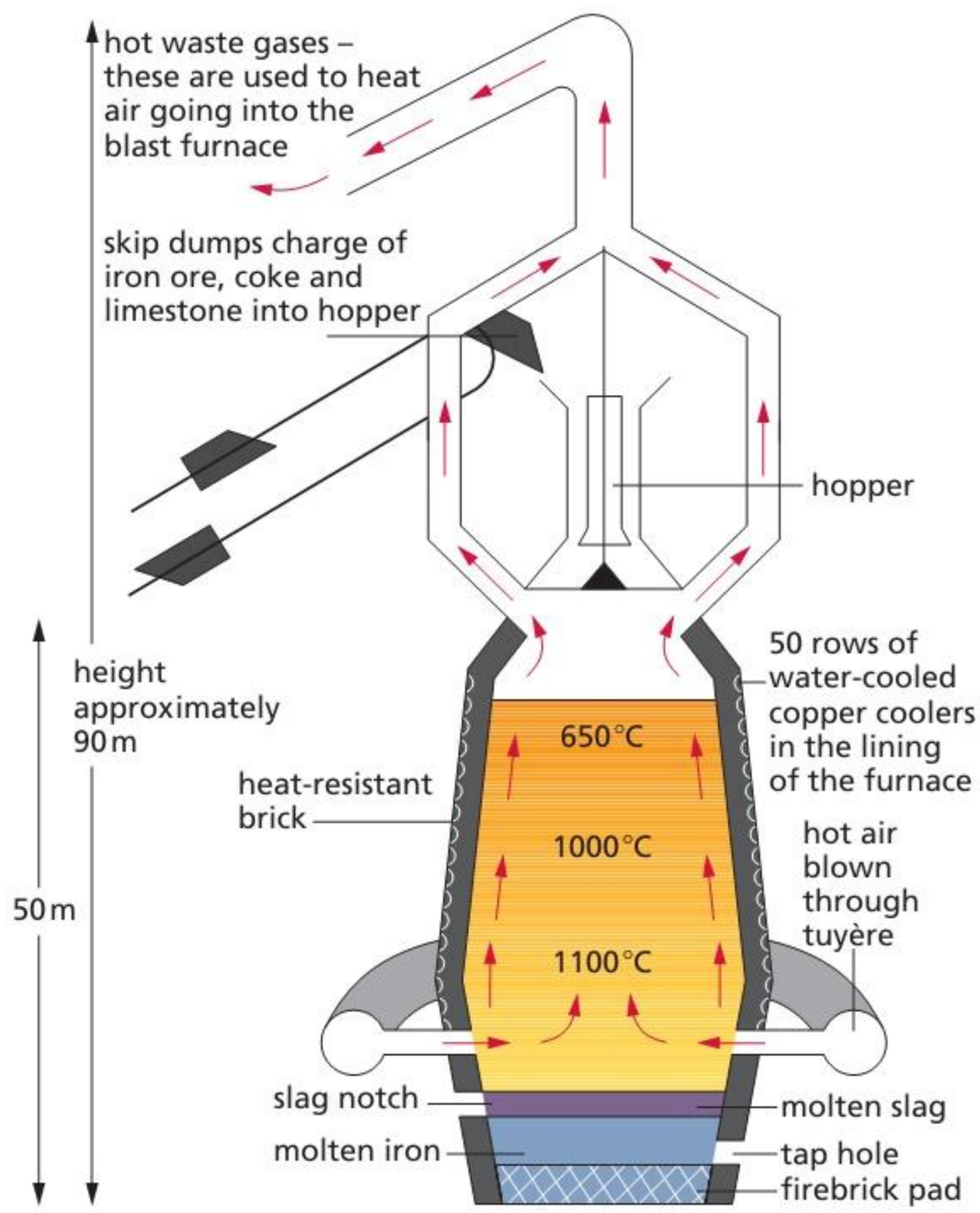


Figure 10.13 Cross-section of a blast furnace.

A number of chemical reactions then follow.

- The limestone begins to decompose:

calcium carbonate	→	calcium oxide	+	carbon dioxide
$CaCO_3(s)$	→	$CaO(s)$	+	$CO_2(g)$

- The carbon dioxide gas produced reacts with more hot coke higher up in the furnace, producing carbon monoxide in an endothermic reaction.

carbon dioxide	+	coke	→	carbon monoxide
$CO_2(g)$	+	$C(s)$	→	$2CO(g)$

- Carbon monoxide is a reducing agent (Chapter 2, p. 14). It rises up the furnace and reduces the iron(III) oxide ore. This takes place at a temperature of around 700 °C:

iron(III) oxide	+	carbon monoxide	→	iron	+	carbon dioxide
$Fe_2O_3(s)$	+	$3CO(g)$	→	$2Fe(l)$	+	$3CO_2(g)$

  The molten iron trickles to the bottom.
- Calcium oxide is a base and this reacts with acidic impurities such as silicon(IV) oxide in the iron, to form a slag which is mainly calcium silicate.

calcium oxide	+	silicon(IV) oxide	→	calcium silicate
$CaO(s)$	+	$SiO_2(s)$	→	$CaSiO_3(l)$

  The slag trickles to the bottom of the furnace, but because it is less dense than the molten iron, it floats on top of it.

Generally, metallic oxides, such as calcium oxide (CaO), are basic and non-metallic oxides, such as silicon(IV) oxide (SiO_2), are acidic.

Certain oxides, such as carbon monoxide (CO), are neutral and others, such as zinc oxide (ZnO), are amphoteric (p. 156).

The molten iron, as well as the molten slag, may be **tapped off** (run off) at regular intervals.

The waste gases, mainly nitrogen and oxides of carbon, escape from the top of the furnace. They are used in a heat exchange process to heat incoming air and so help to reduce the energy costs of the process. Slag is the other waste material. It is used by builders and road makers (Figure 10.14) for foundations.

Figure 10.14 Slag is used in road foundations.

The extraction of iron is a continuous process and is much cheaper to run than an electrolytic method.

The iron obtained by this process is known as 'pig' or cast iron and contains about 4% carbon (as well as some other impurities). The name pig iron arises from the fact that if it is not subsequently converted into steel it is poured into moulds called pigs. Because of its brittle and hard nature, the iron produced by this process has limited use. Gas cylinders are sometimes made of cast iron, since they are unlikely to get deformed during their use.

The majority of the iron produced in the blast furnace is converted into different steel alloys (p. 166) such as manganese and tungsten steels as well as the well-known example of stainless steel (p. 166).

The annual production of iron worldwide is 1536 million tonnes. The larger blast furnaces are capable of producing 10 000 tonnes of iron per day (Figure 10.15).

Figure 10.15 Redcar steelworks at Teesside, UK.

Zinc extraction

The principal source of zinc is zinc sulfide or zinc blende. This ore occurs mainly in Australia, Canada and the US. The zinc ore is first concentrated by a process called **froth flotation**. The crushed ore is fed into tanks of water containing a chemical frothing agent. Air is blown through so that the whole mixture froths up. The rock particles become soaked with water and sink to the bottom of the tank. However, the zinc sulfide particles, into which the water cannot soak, are carried to the top of the tank by the air bubbles and are skimmed off and dried. This ore now contains 55–75% of zinc sulfide. The zinc sulfide is then heated very strongly in a current of air in a furnace (Figure 10.16) to convert it to the oxide:

zinc sulfide	+ oxygen	→	zinc oxide	+ sulfur dioxide
$2ZnS(s)$	$+ 3O_2(g)$	$\rightarrow$	$2ZnO(s)$	$+ 2SO_2(g)$

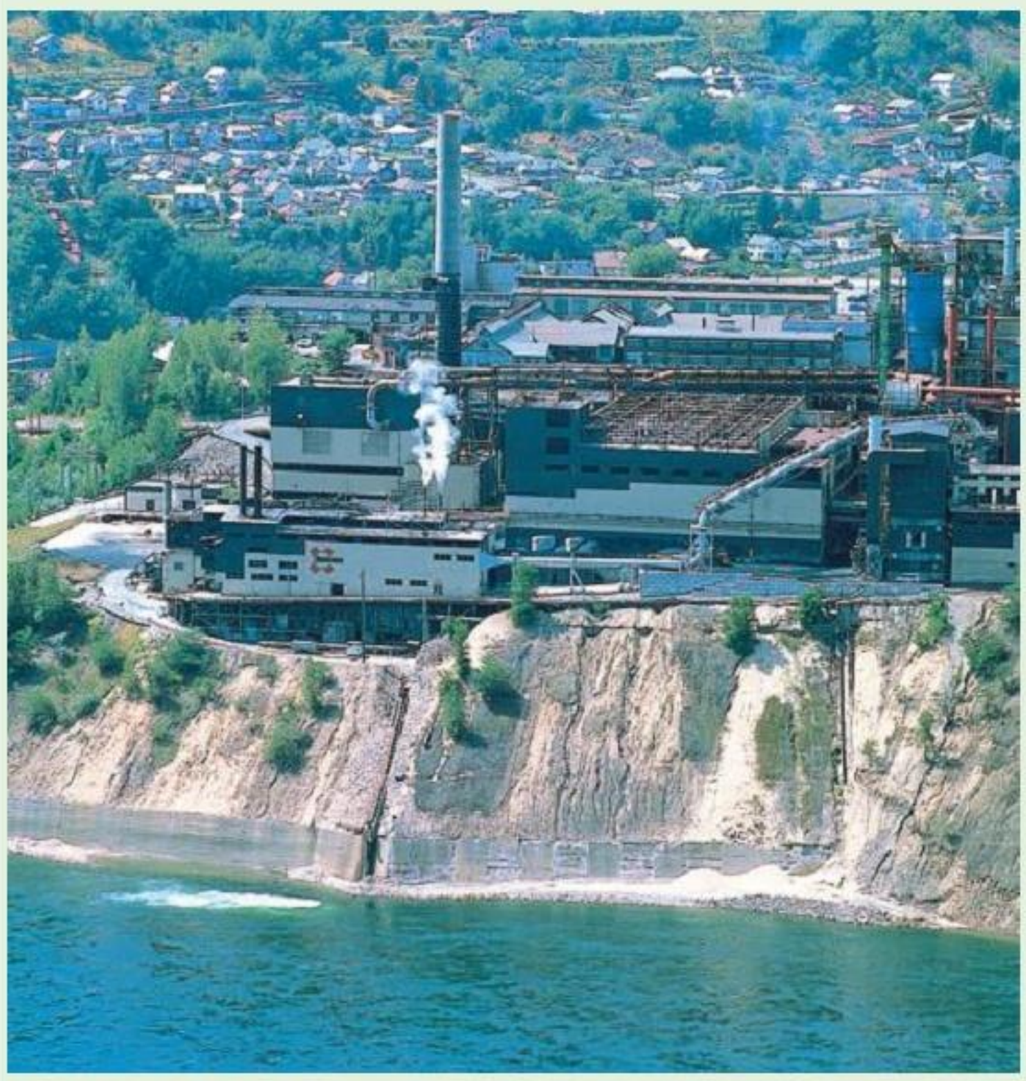


Figure 10.16 Zinc is extracted in a furnace like this one.

The sulfur dioxide is a useful co-product and is used in the manufacture of sulfuric acid. The zinc oxide is mixed with powdered coke in a furnace and heated very strongly to a temperature of approximately 1400 °C. The zinc oxide is reduced by the coke to zinc:

zinc oxide	+ coke (carbon)	→	zinc	+ carbon monoxide
$ZnO(s)$	$+ C(s)$	$\rightarrow$	$Zn(g)$	$+ CO(g)$

The mixture of zinc vapour and carbon monoxide passes through an outlet near the top of the furnace and the zinc metal cools and condenses. The heating costs of the furnace are reduced by burning the carbon monoxide which is produced.

Zinc is used in alloys such as brass. It is also used to galvanise steel and for electrodes in batteries.

Extraction of unreactive metals

Silver and gold

These are very unreactive metals. Silver exists mainly as silver sulfide, Ag_2S (silver glance). The extraction involves treatment of the pulverised ore with sodium cyanide. Zinc is then added to displace the silver from solution. The pure metal is obtained by electrolysis. Silver also exists to a small extent native in the Earth's crust. Gold is nearly always found in its native form (Figure 10.17). It is also obtained in significant amounts during both the electrolytic refining of copper and the extraction of lead.

Figure 10.17 These people are panning for gold.

Silver and gold, because of their resistance to corrosion, are used to make jewellery. Both of these metals are also used in the electronics industry because of their high electrical conductivity.

Questions

1 How does the method used for extracting a metal from its ore depend on the metal's position in the reactivity series?
2 'It is true to say that almost all the reactions by which a metal is extracted from its ore are reduction reactions.' Discuss this statement with respect to the extraction of iron, aluminium and zinc.
3 Titanium is extracted from rutile (TiO_2). Use your research skills, including the Internet, to suggest the detail of the extraction process.
4 Write a word and symbol equation for the reduction of Cu(I) sulfide to copper by heating in air.

Metal waste

Recycling has become commonplace in recent years (Figure 10.18). Why should we really want to recycle metals? Certainly, if we extract fewer metals from the Earth then the existing reserves will last that much longer. Also, recycling metals prevents the creation of a huge environmental problem (Figure 10.19). However, one of the main considerations is that it saves money.

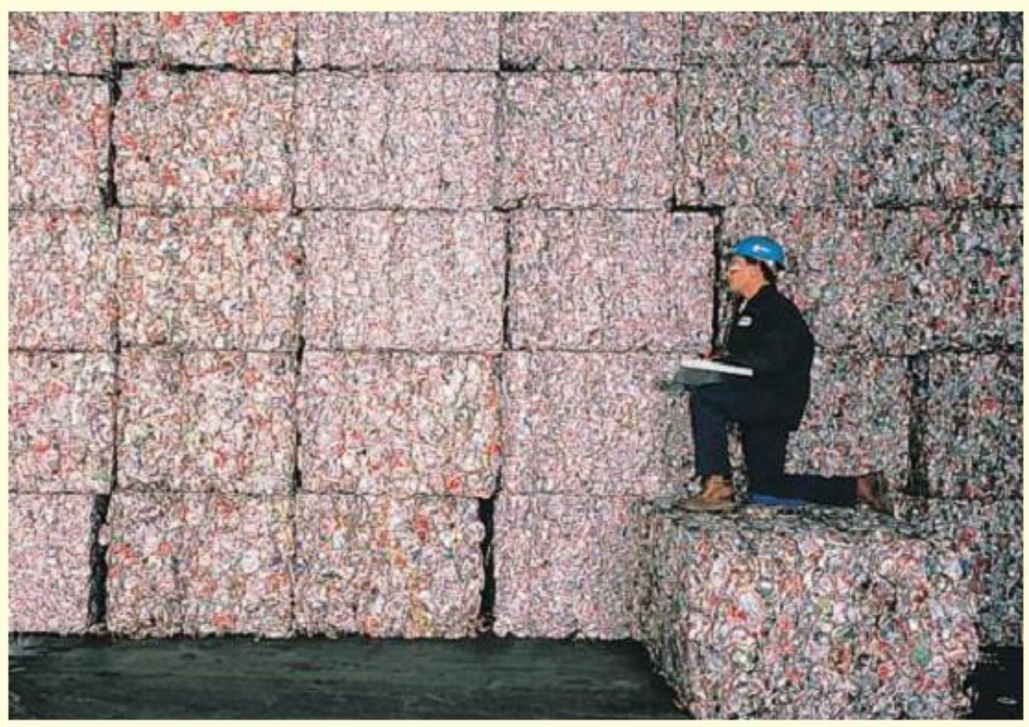


Figure 10.18 Aluminium can recycling.

The main metals which are recycled include aluminium and iron. Aluminium is saved by many households as drinks cans and milk bottle tops, to be melted down and recast. Iron is collected at community tips in the form of discarded household goods and it also forms a large part of the materials collected by scrap metal dealers. Iron is recycled to steel. Many steel-making furnaces run mainly on scrap iron.

Figure 10.19 If we did not recycle metals, then this sight would be commonplace.

Aluminium is especially easy to recycle at low cost. Recycling uses only 5% of the energy needed to extract the metal by electrolysis from bauxite. Approximately 60% of the European need for aluminium is obtained by recycling.

Rusting of iron

After a period of time, objects made of iron or steel will become coated with rust. The rusting of iron is a serious problem and wastes enormous amounts of money each year. Estimates are difficult to make, but it is thought that upwards of £1 billion a year is spent worldwide on replacing iron and steel structures.

Rust is an orange–red powder consisting mainly of hydrated iron(III) oxide ($Fe_2O_3.xH_2O$). Both water and oxygen are essential for iron to rust, and if one of these two substances is not present then rusting will not take place. The rusting of iron is encouraged by salt. Figure 10.20 (p. 162) shows an experiment to show that oxygen (from the air) and water are needed for iron to rust.

Rust prevention

To prevent iron rusting, it is necessary to stop oxygen (from the air) and water coming into contact with it. There are several ways of doing this.

Painting

Ships, lorries, cars, bridges and many other iron and steel structures are painted to prevent rusting (Figure 10.21). However, if the paint is scratched,

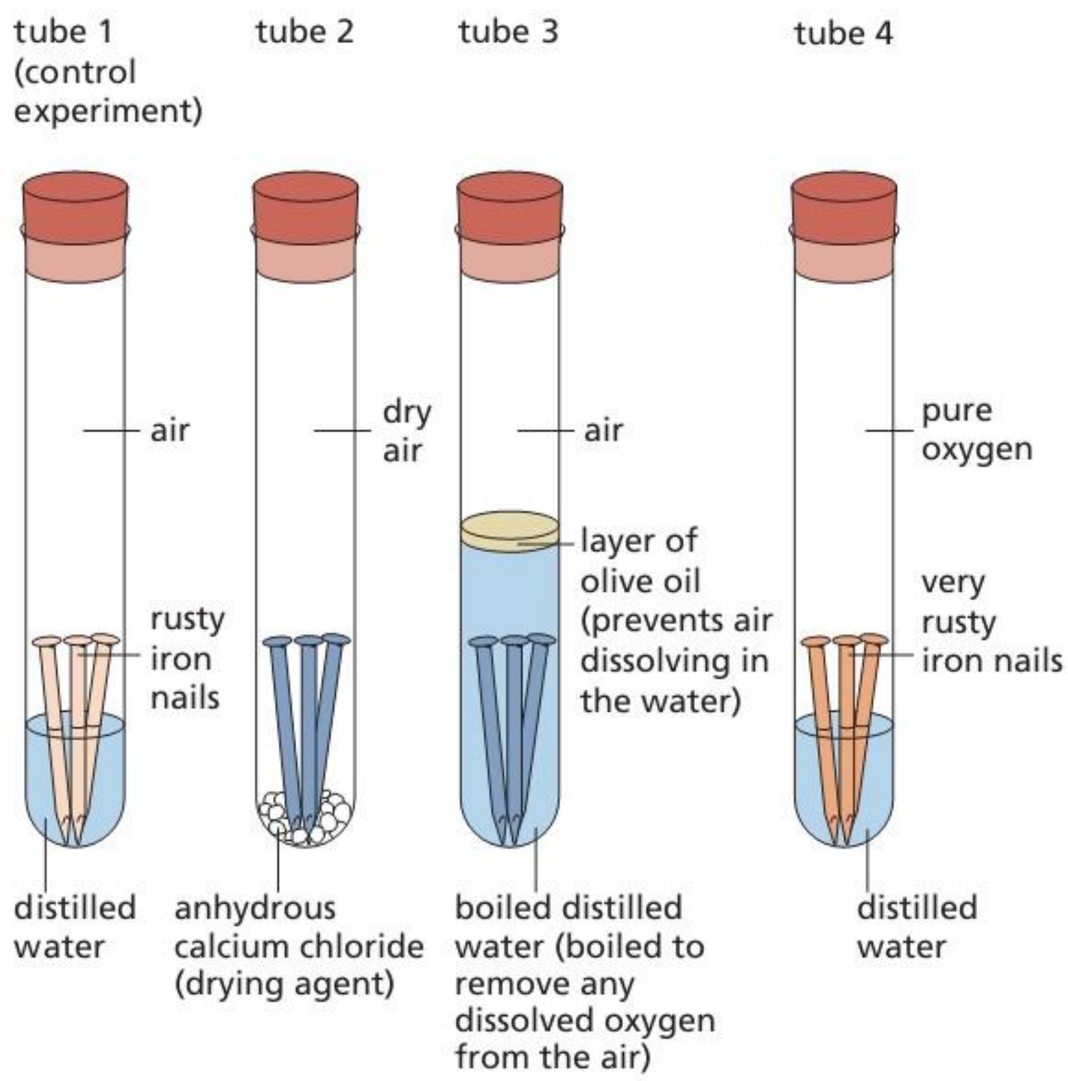


Figure 10.20 Rusting experiment with nails.

Figure 10.21 Painting keeps the air and water away from the steel used to build a ship.

the iron beneath it will start to rust (Figure 10.22) and corrosion can then spread under the paintwork which is still sound. This is why it is essential that the paint is kept in good condition and checked regularly.

Figure 10.22 A brand new car is protected against corrosion (top). However, if the paintwork is damaged, then rusting will result.

Oiling/greasing

The iron and steel in the moving parts of machinery are coated with oil to prevent them from coming into contact with air or moisture. This is the most common way of protecting moving parts of machinery, but the protective film must be renewed.

Coating with plastic

The exteriors of refrigerators, freezers and many other items are coated with plastic, such as PVC, to prevent the steel structure rusting (Figure 10.23).

Plating

Cans for food can be made from steel coated with tin. The tin is deposited on to the steel used to make food cans by dipping the steel into molten tin.

Figure 10.23 A coating of plastic stops metal objects coming into contact with oxygen or water.

Some car bumpers, as well as bicycle handlebars, are electroplated with chromium to prevent rusting. The chromium gives a decorative finish as well as protecting the steel beneath.

Galvanising

Some steel girders, used in the construction of bridges and buildings, are **galvanised**. Coal bunkers and steel dustbins are also galvanised. This involves dipping the object into molten zinc (Figure 10.24). The thin layer of the more reactive zinc metal coating the steel object slowly corrodes and loses electrons to the iron, thereby protecting it. This process continues even when much of the layer of zinc has been scratched away, so the iron continues to be protected.

Figure 10.24 The Burnley Singing Ringing Tree is a sculpture made from galvanised tubes.

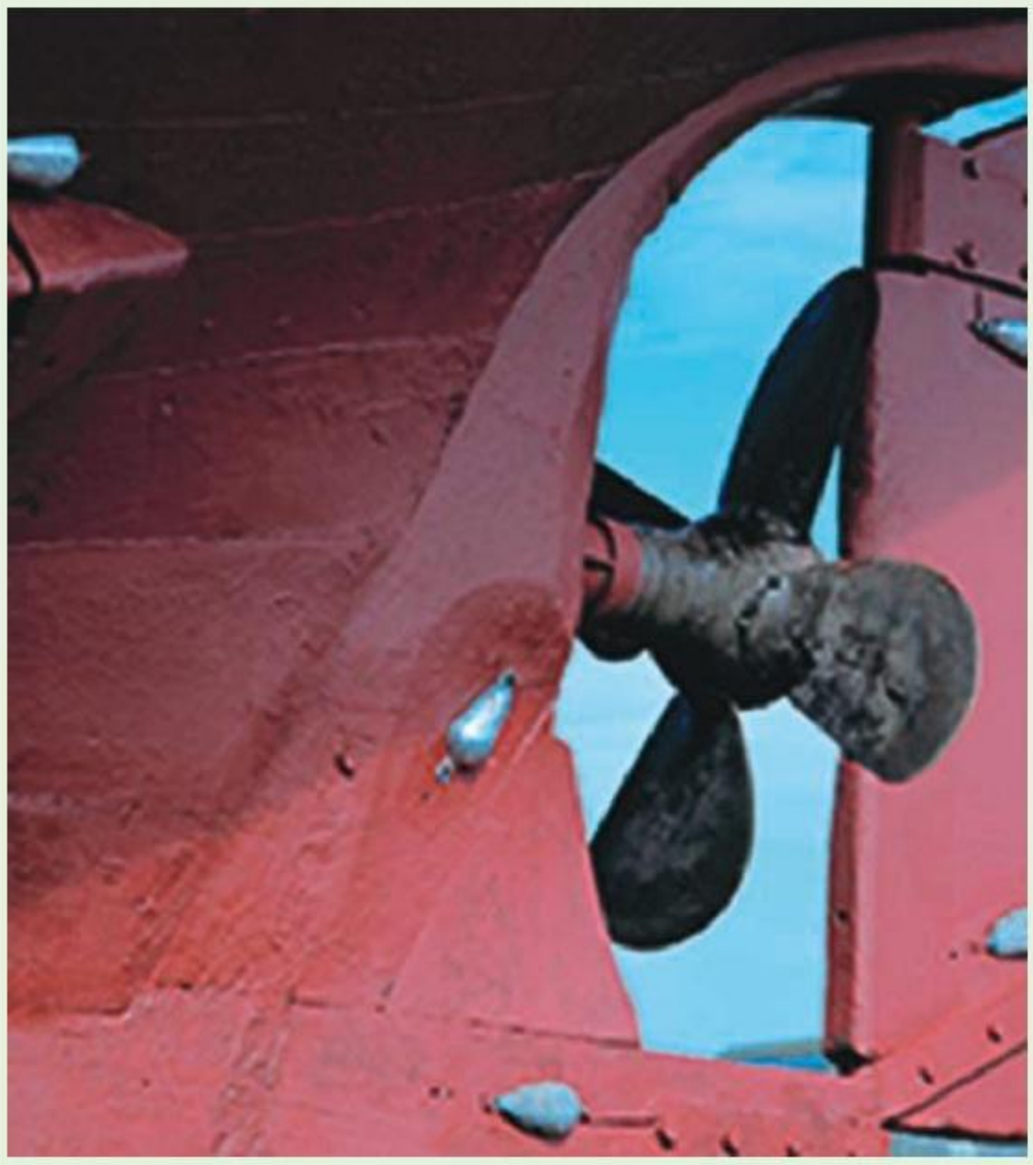
Sacrificial protection

Bars of zinc are attached to the hulls of ships and to oil rigs (as shown in Figure 10.25a). Zinc is above iron in the reactivity series and will react in preference to it and so is corroded. It forms positive ions more readily than the iron:

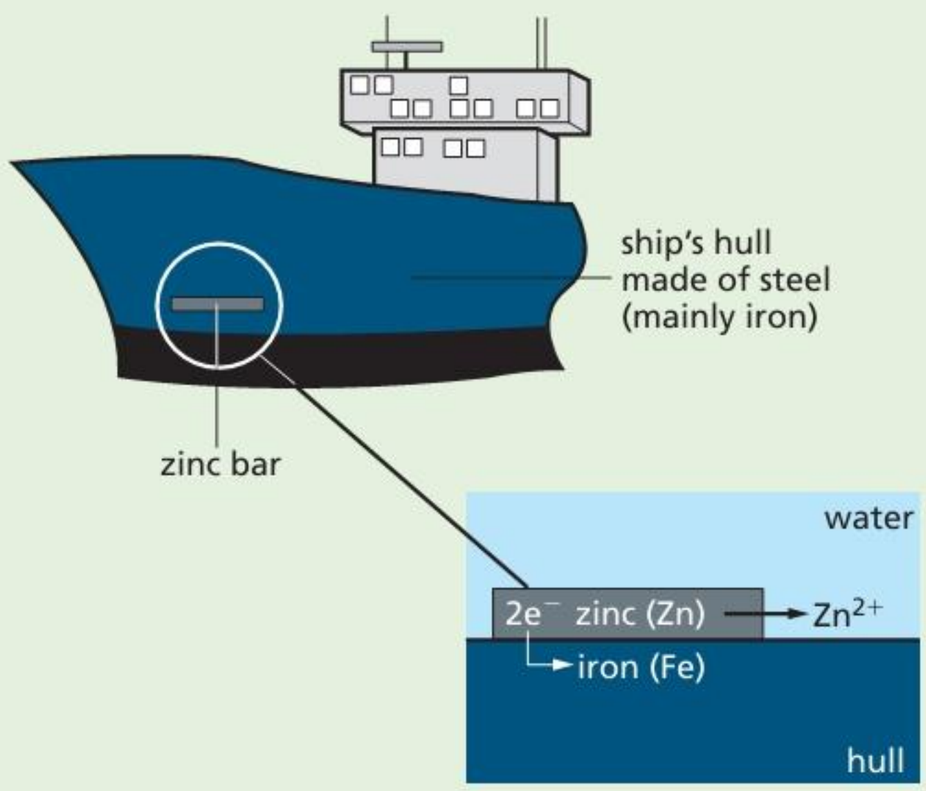
$$Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s)$$

As long as some of the zinc bars remain in contact with the iron structure, the structure will be protected from rusting. When the zinc runs out,

it must be renewed. Gas and water pipes made of iron and steel are connected by a wire to blocks of magnesium to obtain the same result. In both cases, as the more reactive metal corrodes it loses electrons to the iron and so protects it (Figure 10.25b).



a Bars of zinc on the hull of a ship.



b The zinc is sacrificed to protect the steel. Electrons released from the dissolving zinc cause reduction to occur at the surface of the hull.

Figure 10.25 Sacrificial protection.

Corrosion

Rusting is the most common form of corrosion but this term is confined to iron and steel. **Corrosion** is the general name given to the process which takes place when metals and alloys are chemically attacked by oxygen, water or any other substances found in their immediate environment. The metals in the reactivity series will corrode to a greater or lesser extent. Generally, the higher the metal is in the reactivity series, the more rapidly it will corrode. If sodium and potassium were not stored under oil they would corrode very rapidly indeed. Magnesium, calcium and aluminium are usually covered by a thin coating of oxide after initial reaction with oxygen in the air. Freshly produced copper is pink in colour (Figure 5.16a on p. 81). However, on exposure to air, it soon turns brown due to the formation of copper(II) oxide on the surface of the metal.

In more exposed environments, copper roofs and pipes quickly become covered in verdigris. Verdigris is green in colour (Figure 10.26) and is composed of copper salts formed on copper. The composition of verdigris varies depending on the atmospheric conditions, but includes mixed copper(II) carbonate and copper(II) hydroxide ($CuCO_3.Cu(OH)_2$).

Gold and platinum are unreactive and do not corrode, even after thousands of years.

Figure 10.26 Verdigris soon covers copper roofs in exposed environments.

Questions

1 What is rust? Explain how rust forms on structures made of iron or steel.
2 Rusting is a redox reaction. Explain the process of rusting in terms of oxidation and reduction (Chapter 5, p. 73).
3 Design an experiment to help you decide whether steel rusts faster than iron.
4 Why do car exhausts rust faster than other structures made of steel?

Alloys

The majority of the metallic substances used today are **alloys**. Alloys are mixtures of two or more metals and are formed by mixing molten metals thoroughly. It is generally found that alloying produces a metallic substance that has more useful properties than the original pure metals it was made from. For example, the alloy brass is made from copper and zinc. The alloy is harder and more corrosion resistant than either of the metals it is made from.

Steel, which is a mixture of the metal iron and the non-metal carbon, is also considered to be an alloy. Of all the alloys we use, steel is perhaps the most important. Many steels have been produced; they contain not only iron but also carbon and other metals. For example, nickel and chromium are the added metals when stainless steel is produced (Figure 10.27). The chromium prevents the steel from rusting while the nickel makes it harder.

Figure 10.27 A stainless steel exhaust system. Why do you think more people are buying these exhaust systems?

Production of steel

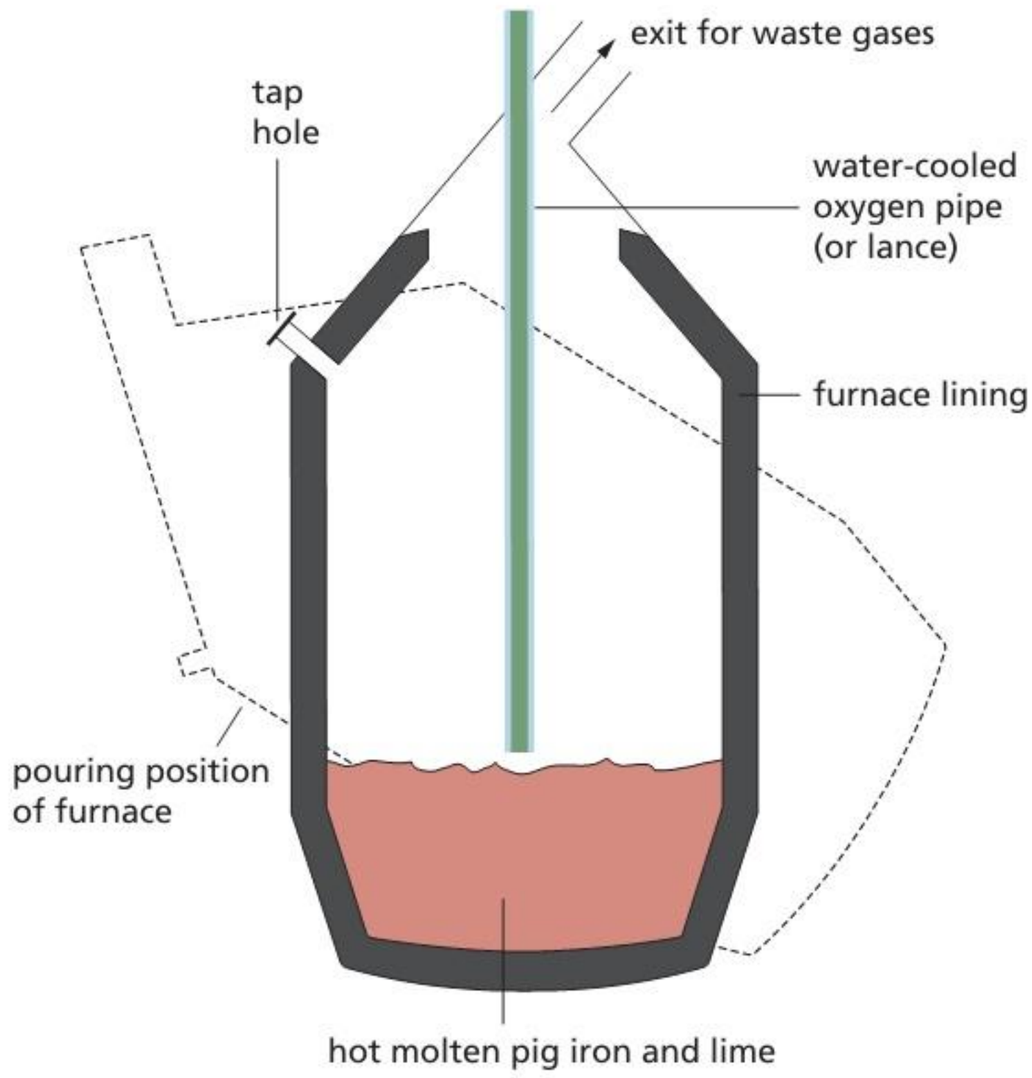
The 'pig iron' obtained from the blast furnace contains between 3% and 5% of carbon and other impurities, such as sulfur, silicon and phosphorus. These impurities make the iron hard and brittle. In order to improve the quality of the metal, most of the impurities must be removed and in doing this, steel is produced.

The impurities are removed in the **basic oxygen process**, which is the most important of the steel-making processes. In this process:

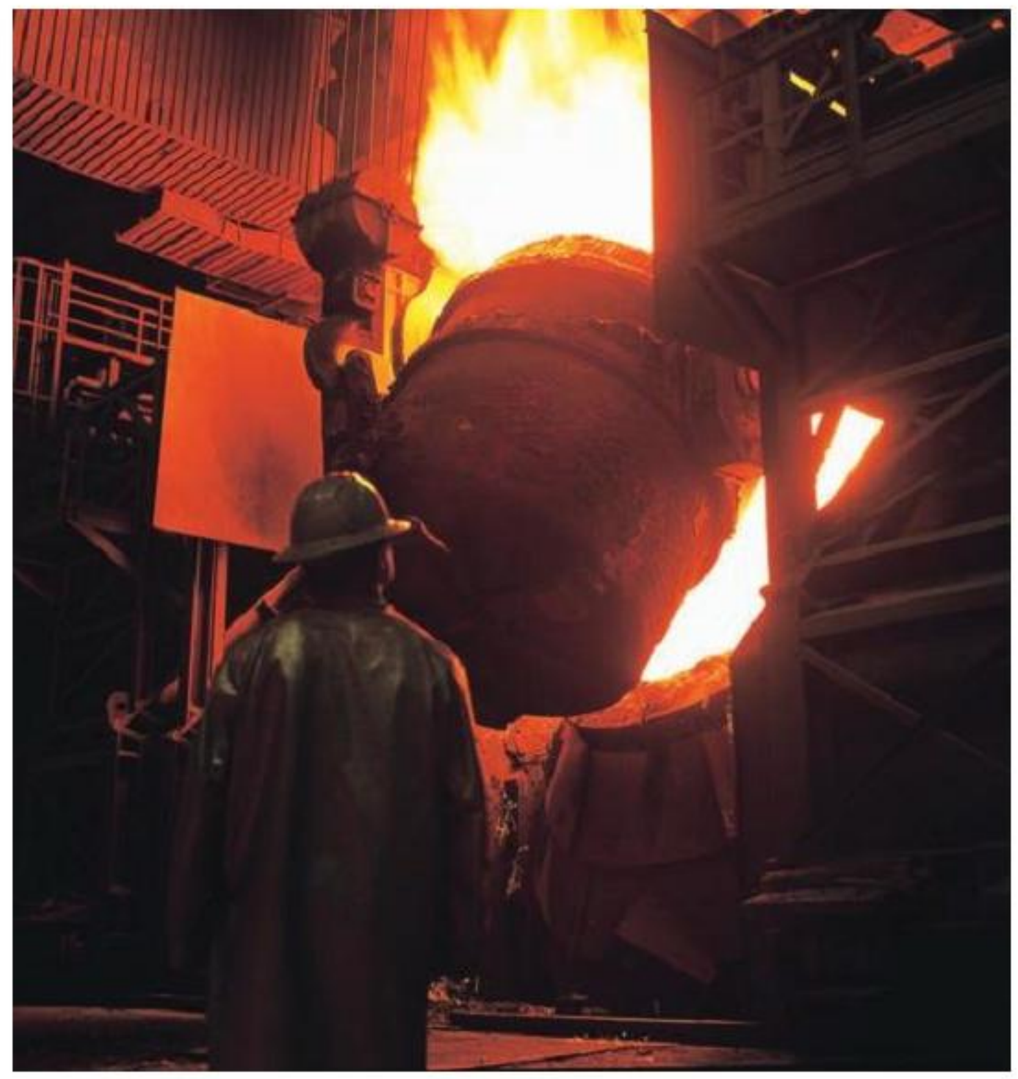
- Molten pig iron from the blast furnace is poured into the basic oxygen furnace (Figure 10.28 p. 166).
- A water-cooled 'lance' is introduced into the furnace and oxygen at 5–15 atm pressure is blown onto the surface of the molten metal.
- Carbon is oxidised to carbon monoxide and carbon dioxide, while sulfur is oxidised to sulfur dioxide. These escape as gases.
- Silicon and phosphorus are oxidised to silicon(IV) oxide and phosphorus pentoxide, which are solid oxides.
- Some calcium oxide (lime) is added to remove these solid oxides as slag. The slag may be skimmed or poured off the surface.
- Samples are continuously taken and checked for carbon content. When the required amount of carbon has been reached, the blast of oxygen is turned off.

The basic oxygen furnace can convert up to 300 tonnes of pig iron to steel per hour. Worldwide production by this process is 430 million tonnes.

There are various types of steel that differ only in their carbon content. The differing amounts of carbon present confer different properties on the steel and they are used for different purposes (Table 10.7). If other types of steel are required then up to 30% scrap steel is added, along with other metals (such as tungsten), and the carbon is burned off.



a A section through the basic oxygen furnace.



b Basic oxygen furnace.

Figure 10.28

Table 10.7 Different types of steel.

Steel	Typical composition	Properties	Uses
Mild steel	99.5% iron, 0.5% carbon	Easily worked Lost most of brittleness	Car bodies, large structures, machinery
Hard steel	99% iron, 1% carbon	Tough and brittle	Cutting tools, chisels, razor blades
Manganese steel	87% iron, 13% manganese	Tough, springy	Drill bits, springs
Stainless steel	74% iron, 18% chromium, 8% nickel	Tough, does not corrode	Cutlery, kitchen sinks, surgical instruments, chemical plant reaction vessels
Tungsten steel	95% iron, 5% tungsten	Tough, hard, even at high temperatures	Edges of high-speed cutting tools

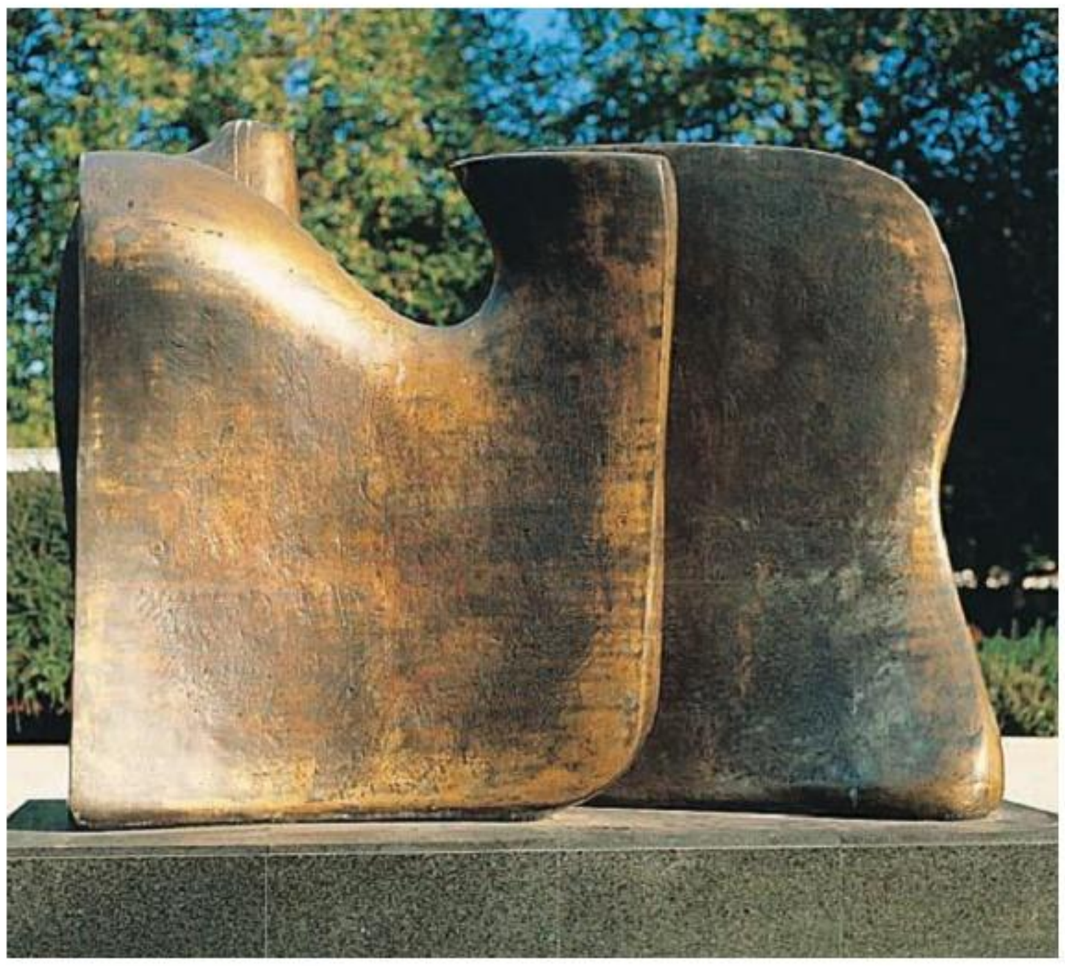
Steel recycling

The recycling of scrap steel contributes 310 million tonnes to the world supply of the alloy – 750 million tonnes. It has been calculated that the energy savings are equivalent to 160 million tonnes of coal. Also, it has been calculated that the raw materials conserved are equivalent to 200 million tonnes of iron ore.

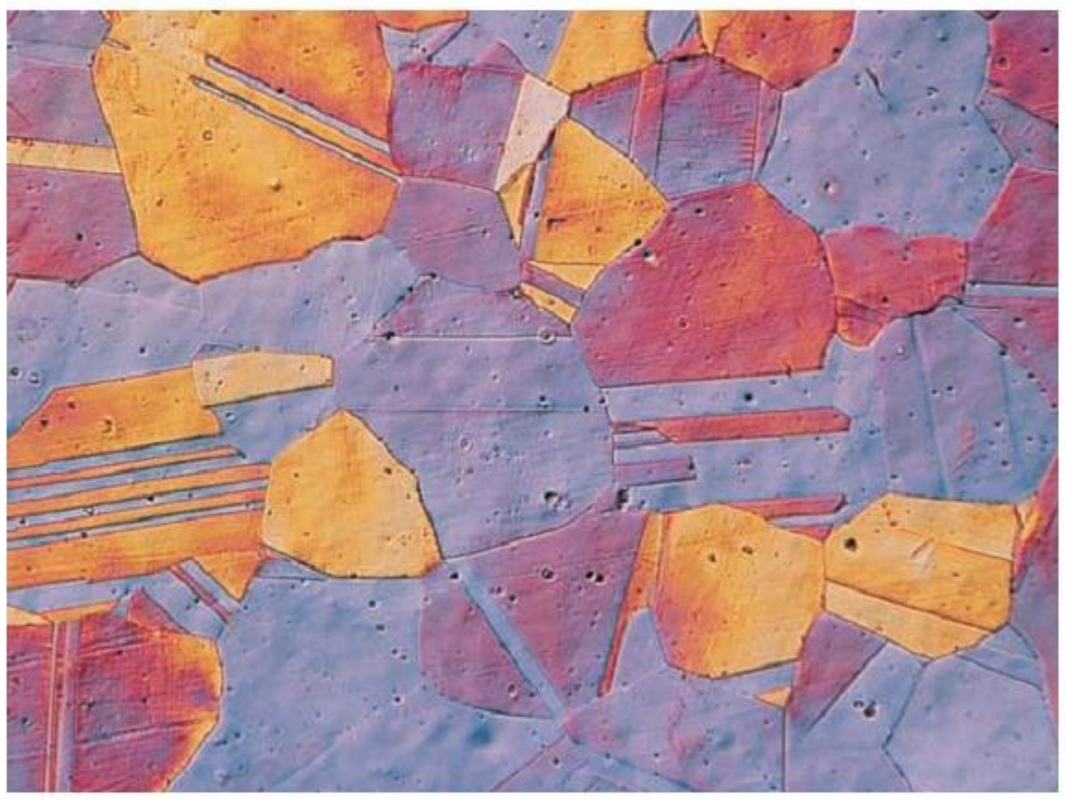
Alloys to order

Just as the properties of iron can be changed by alloying, so the same can be done with other useful metals. Metallurgists have designed alloys to suit a wide variety of different uses. Many thousands of alloys are now made, with the majority being 'tailor-made' to do a particular job (Figure 10.29).

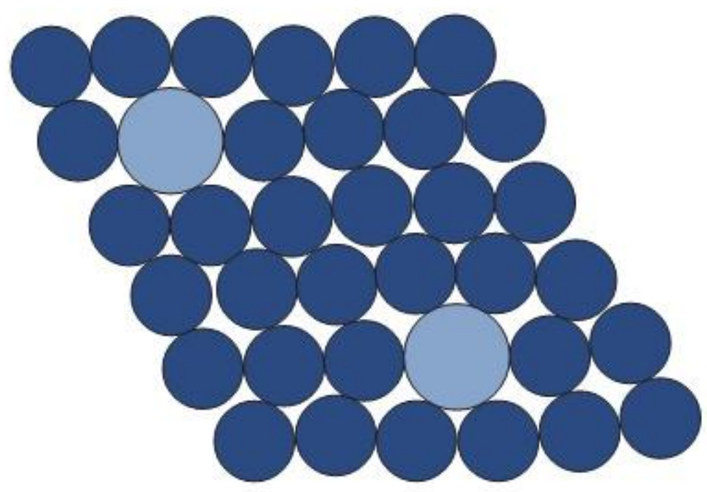
Table 10.8 shows some of the more common alloys, together with some of their uses.



a Bronze is often used in sculptures.



b A polarised light micrograph of brass showing the distinct grain structure of this alloy.



c Alloy structure. The dark circles represent atoms of a metal; the pale circles are the larger atoms of a different metal added to make the alloy. The different size of these atoms gives the alloy different physical properties from those of the pure metal.

Figure 10.29

Table 10.8 Uses of common alloys.

Alloy	Composition	Use
Brass	65% copper, 35% zinc	Jewellery, machine bearings, electrical connections, door furniture
Bronze	90% copper, 10% tin	Castings, machine parts
Cupro-nickel	30% copper, 70% nickel	Turbine blades
	75% copper, 25% nickel	Coinage metal
Duralumin	95% aluminium, 4% copper, 1% magnesium, manganese and iron	Aircraft construction, bicycle parts
Magnalium	70% aluminium, 30% magnesium	Aircraft construction
Pewter	30% lead, 70% tin, a small amount of antimony	Plates, ornaments and drinking mugs
Solder	70% lead, 30% tin	Connecting electrical wiring

Questions

1 Calcium oxide is a base. It combines with solid, acidic oxides in the basic oxygen furnace.
Write a chemical equation for one of these oxides reacting with the added lime.

2 'Many metals are more useful to us when mixed with some other elements.' Discuss this statement with respect to stainless steel.

Checklist

After studying Chapter 10 you should know and understand the following terms.

- **Alloy** Generally, a mixture of two or more metals (for example, brass is an alloy of zinc and copper) or of a metal and a non-metal (for example, steel is an alloy of iron and carbon, sometimes with other metals included). They are formed by mixing the molten substances thoroughly. Generally, it is found that alloying produces a metallic substance which has more useful properties than the original pure metals it was made from.
- **Amphoteric hydroxide** A hydroxide which can behave as an acid (react with an alkali) or a base (react with an acid), for example zinc hydroxide.
- **Blast furnace** A furnace for smelting iron ores such as haematite (Fe_2O_3) and magnetite (Fe_3O_4) to produce pig (or cast) iron. In a modified form it can be used to extract metals such as zinc.
- **Competition reactions** Reactions in which metals compete for oxygen or anions. The more reactive metal:
  - takes the oxygen from the oxide of a less reactive metal
  - displaces the less reactive metal from a solution of that metal salt – this type of competition reaction is known as a displacement reaction.
- **Corrosion** The process that takes place when metals and alloys are chemically attacked by oxygen, water or any other substances found in their immediate environment.
- **Metal extraction** The method used to extract a metal from its ore depends on the position of the metal in the reactivity series.
  - Reactive metals are usually difficult to extract. The preferred method is by electrolysis of the molten ore (electrolytic reduction); for example, sodium from molten sodium chloride.
  - Moderately reactive metals (those near the middle of the reactivity series) are extracted using a chemical reducing agent (for example carbon) in a furnace; for example, iron from haematite in the blast furnace.
  - Unreactive metals, for example gold and silver, occur in an uncombined (native) state as the free metal.
- **Metal ion precipitation** The reactions in which certain metal cations form insoluble hydroxides. The colours of these insoluble hydroxides can be used to identify the metal cations which are present; for example, copper(II) hydroxide is a blue precipitate.
- **Ore** A naturally occurring mineral from which a metal can be extracted.
- **Reactivity series of metals** An order of reactivity, giving the most reactive metal first, based on results from experiments with oxygen, water and dilute hydrochloric acid.
- **Recycling metals** Metal drink cans such as those made of aluminium are collected in large 'banks' for the sole purpose of recycling them. Reusing the metal in this way saves money.
- **Rust** A loose, orange–brown, flaky layer of hydrated iron(III) oxide found on the surface of iron or steel. The conditions necessary for rusting to take place are the presence of oxygen and water. The rusting process is encouraged by other substances such as salt. It is an oxidation process.
- **Rust prevention** To prevent iron rusting it is necessary to stop oxygen and water coming into contact with it. The methods employed include painting, oiling/greasing, coating with plastic, plating, galvanising and sacrificial protection.

Metals

Additional questions

1 Use the following list of metals to answer the questions **a** to **i**: iron, calcium, potassium, gold, aluminium, magnesium, sodium, zinc, platinum, titanium.

a Which of the metals is found native?
b Which of the metals is found in nature as the ore:
(i) rock salt? **(ii)** rutile?
c Which metal has a carbonate found in nature called chalk?
d Which of the metals will not react with oxygen to form an oxide?
e Which of the metals will react violently with cold water?
f Choose one of the metals in your answer to **e** and write a balanced chemical equation for the reaction which takes place.
g Which of the metals has a protective oxide coating on its surface?
h Which of the metals reacts very slowly with cold water but extremely vigorously with steam?
i Which of the metals is used to galvanise iron?

2

Copper
↓
heat in air (oxygen)
↓
Black solid **A**
↓
heat with powdered magnesium
↓
Brown–pink solid **B** + white powder **C**
↓
dilute hydrochloric acid
↓
Brown–pink solid **B** + colourless solution **D** + water
↓
filter and evaporate filtrate to dryness
↓
Solid **D**
↓
electrolysis of molten **D**
↓
Silvery metal **E** + green gas **F**

a Name and give the formulae of the substances **A** to **F**.
b Write balanced chemical equations for the reactions in which:
(i) black solid **A** was formed
(ii) white powder **C** and brown–pink solid **B** were formed
(iii) colourless solution **D** was formed.
c The reaction between black solid **A** and magnesium is a redox reaction. With reference to this reaction, explain what you understand by this statement.
d Write anode and cathode reactions for the processes which take place during the electrolysis of molten **D**.
e Suggest a use for:
(i) brown–pink solid **B**
(ii) silvery metal **E**
(iii) green gas **F**.

3 Explain the following:
a metals such as gold and silver occur native in the Earth's crust
b the parts of shipwrecks made of iron rust more slowly in deep water
c zinc bars are attached to the structure of oil rigs to prevent them from rusting
d copper roofs quickly become covered with a green coating when exposed to the atmosphere
e recycling metals can save money.

4 Iron is extracted from its ores haematite and magnetite. Usually it is extracted from haematite (iron(III) oxide). The ore is mixed with limestone and coke and reduced to the metal in a blast furnace. The following is a brief outline of the reactions involved.

coke + oxygen → gas **X**
gas **X** + coke → gas **Y**
iron(III) oxide + gas **Y** → iron + gas **X**

a Name the gases **X** and **Y**.
b Give a chemical test to identify gas **X**.
c Write balanced chemical equations for the reactions shown above.
d The added limestone is involved in the following reactions:

limestone → calcium oxide + gas **X**
calcium oxide + silicon(IV) oxide → slag

(i) Give the chemical names for limestone and slag.
(ii) Write balanced chemical equations for the reactions shown above.

(iii) Why is the reaction between calcium oxide and silicon(IV) oxide called an acid–base reaction?

(iv) Describe what happens to the liquid iron and slag when they reach the bottom of the furnace.

e Why do you think that the furnace used in the extraction of iron is called a blast furnace?

5 The iron obtained from the blast furnace is known as pig or cast iron. Because of the presence of impurities, such as carbon, it has a hard and brittle nature. Most of this type of iron is therefore converted into steel in the basic oxygen process. During this process either all or some of the carbon is removed. Calculated quantities of other elements are then added to produce the required type of steel.

a Explain the meaning of the term alloy as applied to steel.

b Name two impurities, other than carbon, which are present in cast iron and which are removed completely during the steel manufacture.

c Describe the method of steel manufacture used which removes the impurities referred to in **b**.

d Name two metallic elements which may be added to the basic oxygen furnace to produce different varieties of steel.

e Give two uses of stainless steel.

f Give two advantages of stainless steel compared with cast iron.

6 The metal titanium is extracted from its ore rutile (TiO_2).

a Describe the processes involved in the extraction of titanium from rutile.

b Name the area of the Periodic Table in which titanium is found.

c The final part of the process involves the following reaction:

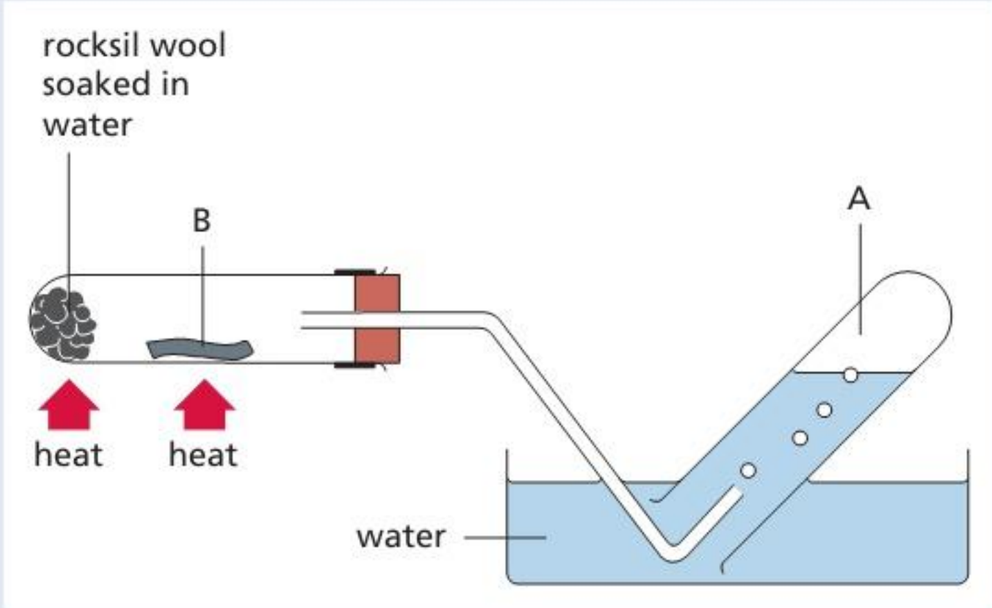
$$TiCl_4(l) + 4Na(l) \rightarrow 4NaCl(s) + Ti(s)$$

(i) Calculate the quantity of titanium produced from 19 tonnes of titanium(IV) chloride ($TiCl_4$).

(ii) How much sodium would have to be used to extract the titanium from the 19 tonnes of $TiCl_4$ used in (**i**)?

d Give two uses of titanium metal.

7 Zinc can be reacted with steam using the apparatus shown. When gas **A** is collected, mixed with air and ignited it gives a small pop. A white solid **B** remains in the test tube when the reaction has stopped and the apparatus cooled down.



a Name and give the formula of gas **A**.

b (i) Name the product formed when gas **A** burns in air.

(ii) Write a balanced chemical equation for this reaction.

c (i) Name white solid **B**.

(ii) Write a balanced chemical equation to represent the reaction between magnesium and steam.

d Name two other metals which could be safely used to replace zinc and produce another sample of gas **A**.

e When zinc reacts with dilute hydrochloric acid, gas **A** is produced again. Write a balanced chemical equation to represent this reaction and name the other product of this reaction.

8 Copper is extracted in the final stage by roasting copper(I) sulfide (Cu_2S) in a current of air. This converts the sulfide to the metal.

$$Cu_2S(s) + O_2(g) \rightarrow 2Cu(s) + SO_2(g)$$

a (i) What is the name of the chemical process by which copper(I) sulfide is converted to copper?

(ii) At what temperature does this process take place?

b It is a very expensive business to maintain the high temperature needed to efficiently convert copper(I) sulfide to copper. How are the heating costs offset?

c Calculate the quantity of copper(I) sulfide required to produce 1270 tonnes of copper. (A_r: O = 16; S = 32, Cu = 63.5)

d For use in electrical wiring, the copper must be 99.99% pure. How is the copper purified to obtain this high purity?

11 Air and water

The air
How did our atmosphere form?
The structure of the atmosphere
The composition of the atmosphere

How do we get the useful gases we need from the air?
Some uses of the gases from the air

Ammonia – an important nitrogen-containing chemical
Making ammonia
Making ammonia in the laboratory
Physical properties of ammonia
Chemical properties of ammonia

Artificial fertilisers
Manufacture of ammonium nitrate
The nitrogen cycle
Problems with fertilisers

Atmospheric pollution

Water
The unique properties of water

The water cycle

Hardness in water
Effect of hard water on soap
Removal of hardness

Water pollution and treatment
Making water fit to drink
The 'iron problem'
Sewage treatment

Checklist

Additional questions

In this chapter we will study the gaseous mixture we know as the **air** (or atmosphere) and the most important and plentiful liquid on this planet, **water**. The importance of air and water is not in doubt. Without either of these, life would not be possible on Earth!

The gases in the air are held in an envelope around the Earth by its gravity. The atmosphere is approximately 100 km thick (Figure 11.1), and about 75% of the mass of the atmosphere is found in the layer nearest the Earth called the **troposphere** (Figure 11.2). Beyond this layer, the atmosphere reaches into space but becomes extremely thin. Nearly all atmospheric water vapour (or moisture) is found in the troposphere, which also contains the liquid water in the oceans, rivers and lakes.

Figure 11.1 The lighter blue shows the extent of the atmosphere around the Earth – approximately 100 km.

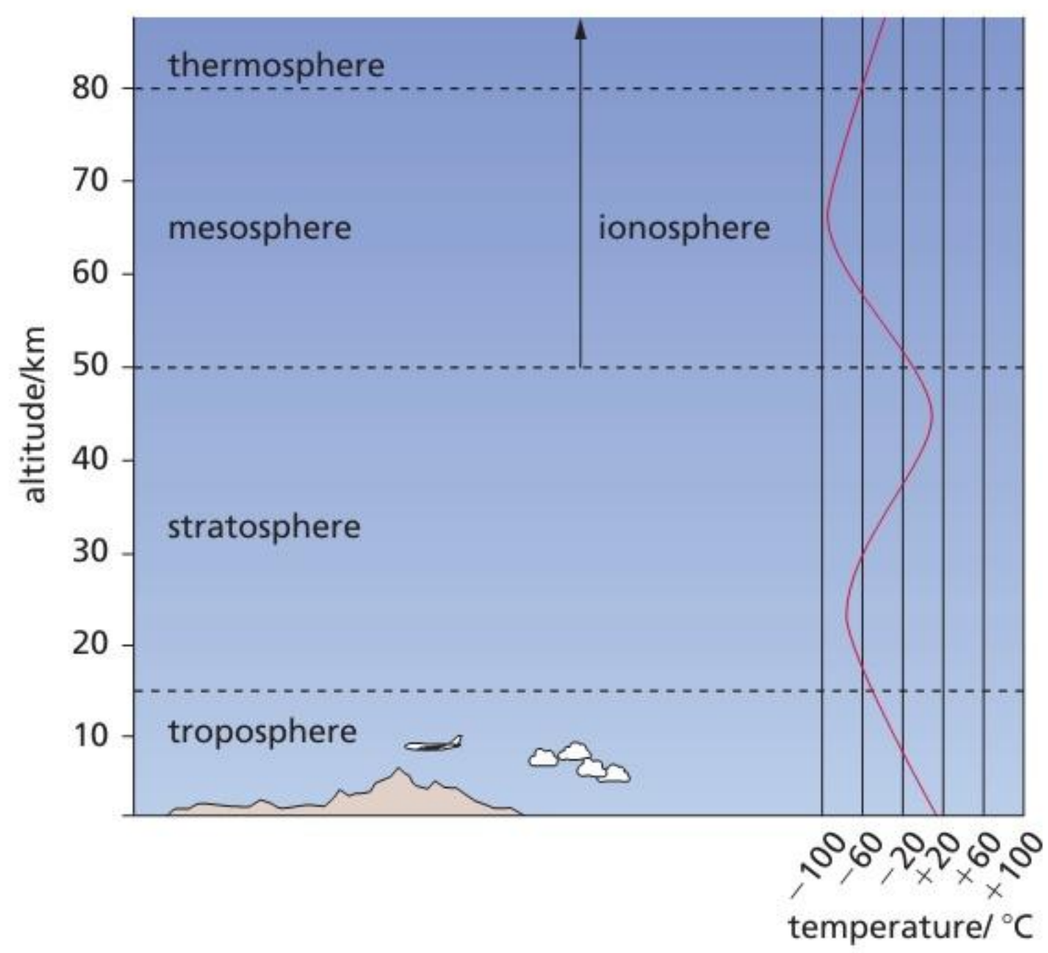


Figure 11.2 The Earth's atmosphere.

The air

How did our atmosphere form?

About 4500 million years ago the Earth was formed. Over a period of time, the Earth cooled to become a molten mass upon which a thin crust formed. Volcanic activity through the crust pushed out huge quantities of gases, such as ammonia, nitrogen, methane, carbon monoxide, carbon dioxide and a small amount of sulfur dioxide, which formed an atmosphere around the Earth (Figure 11.3).

Figure 11.3 Volcanic activity like this expelled gases through the crust to form a secondary atmosphere.

About 3800 million years ago, when the Earth had cooled below 100 °C, the water vapour in this atmosphere condensed and fell as rain. This caused the formation of the first oceans, lakes and seas on the now rapidly cooling Earth. Eventually, early forms of life developed in these oceans, lakes and seas at depths which prevented potentially harmful ultraviolet light from the Sun affecting them.

About 3000 million years ago the first forms of bacteria appeared, followed by algae-like organisms (Figure 11.4).

The process of photosynthesis can be described by the following equation:

$$\text{carbon dioxide} + \text{water} \xrightarrow[\text{chlorophyll}]{\text{sunlight}} \text{glucose} + \text{oxygen}$$

$$6CO_2(g) + 6H_2O(l) \longrightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

Figure 11.4 This algae is related to that early organism produced all those years ago!

These organisms used the light from the Sun to produce their own food, and oxygen was released into the atmosphere as a waste product of **photosynthesis**. This process also acted to reduce the amount of CO_2 in the atmosphere.

The ultraviolet radiation now broke down some of the oxygen molecules in the stratosphere (the level above the troposphere, see below), and some of the oxygen atoms produced formed ozone molecules, $O_3(g)$. Ozone is an unstable molecule which readily decomposes under the action of ultraviolet radiation. However, some of these ozone molecules then re-form.

Ozone is an important gas in the upper atmosphere because it prevents harmful ultraviolet radiation from reaching the Earth. Over many millions of years, the amount of ultraviolet radiation reaching the Earth's surface was reduced significantly.

About 400 million years ago the first land plants appeared on the Earth and so the amount of oxygen, and hence ozone, increased.

Oxygen is a reactive gas and, over millions of years, organisms have adapted to make use of it. The oxygen from the atmosphere was used, along with the carbon they obtained from their food, to produce energy in a process known as **respiration**. The process of respiration can be represented as:

$$\text{glucose} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water} + \text{energy}$$

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l) + \text{energy}$$

The atmosphere as we know it therefore took a very long time to develop to what it is today.

Over recent years, scientists have become aware of a reduction in the amount of ozone in our atmosphere and of the formation of 'ozone holes' in the stratosphere (Figure 11.5). The reduction of ozone in our atmosphere has led to an increased risk of skin cancer as more harmful ultraviolet radiation has reached the surface of the Earth. This is a different type of problem from the greenhouse effect and associated global warming, which is caused mainly by an increase in the amount of carbon dioxide in the atmosphere. For a further discussion of the greenhouse effect and global warming see p. 212.

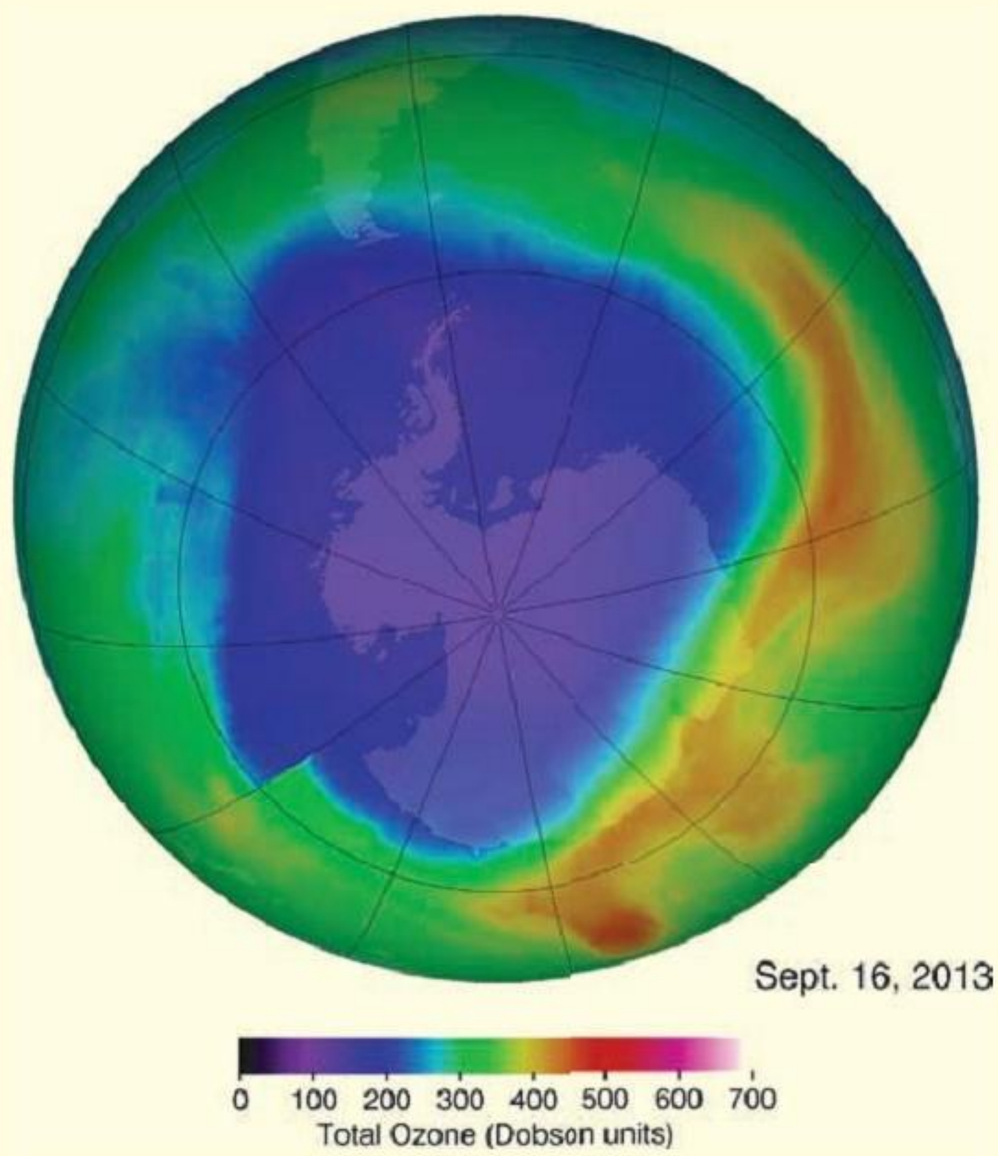


Figure 11.5 This diagram comes from NASA's ozone monitoring programme TOMS (Total Ozone Mapping Spectrometer). The ozone hole over the Antarctic (shown in purple and pink on the diagram) is largest in the Antarctic spring. Note: Dobson Units are a measure of the total amount of ozone in a vertical column from the ground to the top of the atmosphere.

Questions

1 Produce a time line for the formation of the atmosphere, starting with the formation of the Earth 4500 million years ago.
2 Use your research skills to find out:
   a more about the ozone hole problem
   b what precautions are necessary to prevent an increase in skin cancers.

The structure of the atmosphere

The atmosphere is 100 km thick (Figures 11.1 and 11.2) and it is divided into four layers:

- troposphere
- stratosphere
- mesosphere
- thermosphere.

The gases in the atmosphere are held in an envelope around the Earth by its gravity. About 75% of the mass of the atmosphere is found in the layer nearest the Earth called the **troposphere**. Beyond this layer, the atmosphere reaches into space but it becomes extremely thin beyond the mesosphere.

The composition of the atmosphere

If a sample of dry, unpolluted air was taken from any location in the troposphere and analysed, the composition by volume of the sample would be similar to that shown in Table 11.1.

Table 11.1 Composition of the atmosphere.

Component	%
Nitrogen	78.08
Oxygen	20.95
Argon	0.93
Carbon dioxide	0.04
Neon	0.002
Helium	0.0005
Krypton	0.0001
Xenon plus minute amounts of other gases	0.00001

Measuring the percentage of oxygen in the atmosphere

When 100 cm^3 of air is passed backwards and forwards over heated copper turnings it is found that the amount of gas decreases (Figure 11.6). This is because the reactive part of the air, the oxygen gas, is reacting with the copper to form black copper(II) oxide (Figure 11.7). In such an experiment, the volume of gas in the syringe decreases from 100 cm^3 to about 79 cm^3, showing that the air contained 21 cm^3 of oxygen gas. The percentage of oxygen gas in the air is:

$$\frac{21}{100} \times 100 = 21\%$$

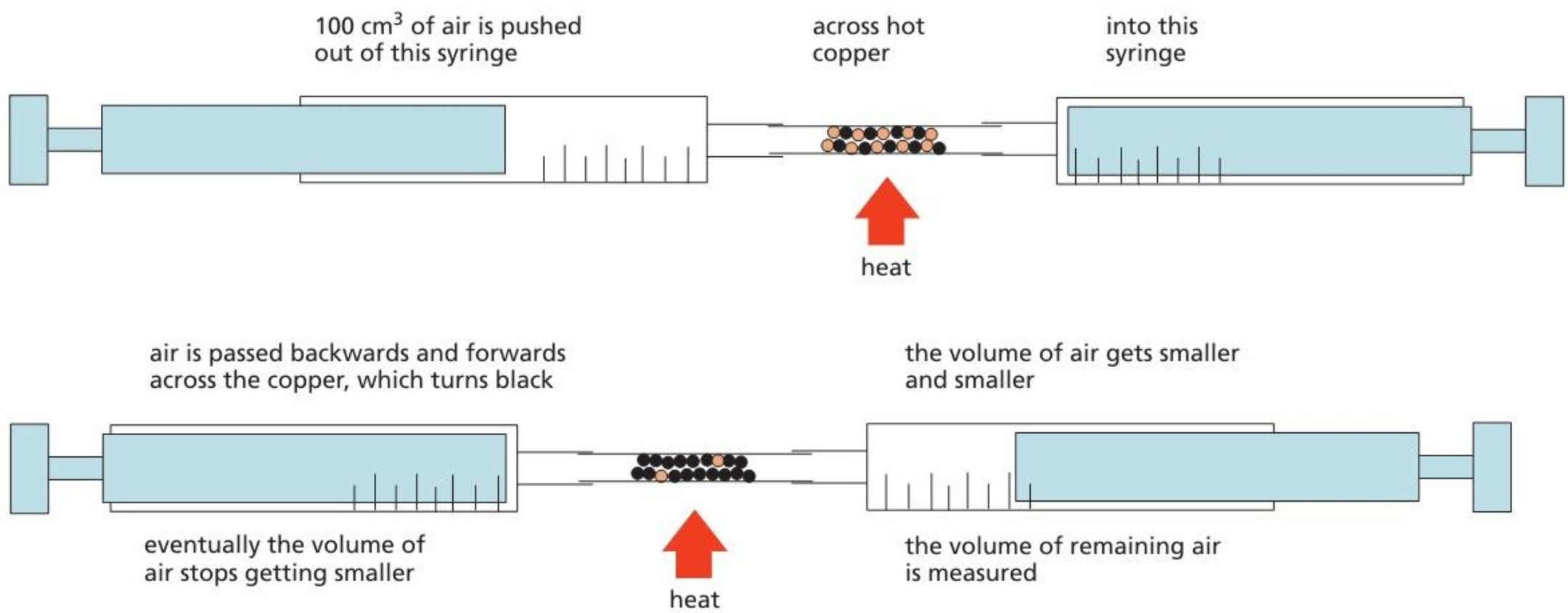


Figure 11.6 This apparatus can be used to find out the volume of oxygen gas in the air.

Figure 11.7 Copper turnings before and after reaction.

The composition of the atmosphere is affected by the following factors:

- respiration
- photosynthesis
- volcanic activity
- radioactive decay, in which helium is formed
- human activity, involving burning of fossil fuels, in which carbon dioxide and water vapour are produced as well as other gases (Chapter 6, p. 92). Human activity of this kind has given rise to an increase in the amount of carbon dioxide in the atmosphere over the last 200 years in particular.

Compare the components of our atmosphere with those of the other planets in the solar system (as shown in Table 11.2).

Table 11.2 Atmospheres of the other planets* in the solar system.

Planet	Atmosphere
Mercury	No atmosphere – the gases were burned off by the heat of the Sun
Venus	Carbon dioxide and sulfur dioxide
Mars	Mainly carbon dioxide
Jupiter	Ammonia, helium, hydrogen, methane
Saturn	Ammonia, helium, hydrogen, methane
Uranus	Ammonia, helium, hydrogen, methane
Neptune	Helium, hydrogen, methane

* Pluto is no longer recognised as a planet.

Questions

1. Draw a pie chart to show the data given in Table 11.1.
2. Is air a compound or a mixture? Explain your answer.
3. Design an experiment to find out how much oxygen there is in exhaled air.

How do we get the useful gases we need from the air?

Air is the major source of oxygen, nitrogen and the noble gases. The gases are obtained by fractional distillation of liquid air but it is a complex process, involving several different steps (Figure 11.8).

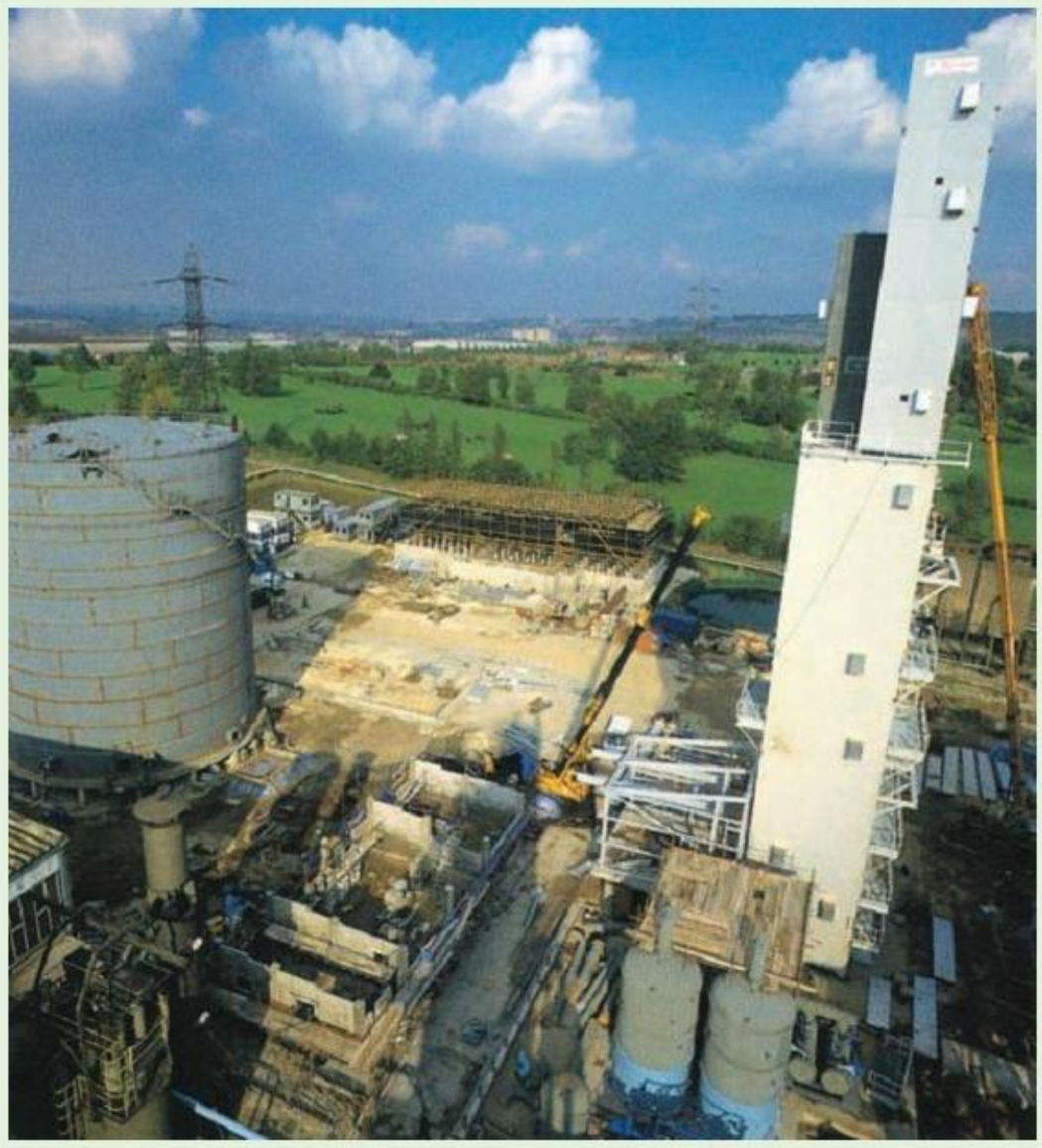


Figure 11.8 Large amounts of gases are obtained from the fractional distillation of liquid air in industrial plants such as this.

- The air is passed through fine filters to remove dust.
- The air is cooled to about –80 °C to remove water vapour and carbon dioxide as solids. If these are not removed, then serious blockages of pipes can result.
- Next, the cold air is compressed to about 100 atm of pressure. This warms up the air, so it is passed into a heat exchanger to cool it down again.
- The cold, compressed air is allowed to expand rapidly, which cools it still further.
- The process of compression followed by expansion is repeated until the air reaches a temperature below –200 °C. At this temperature the majority of the air liquefies (Table 11.3).

Table 11.3 Boiling points of atmospheric gases.

Gas	Boiling point/°C
Helium	–269
Neon	–246
Nitrogen	–196
Argon	–186
Oxygen	–183
Krypton	–157
Xenon	–108

- The liquid air is passed into a fractionating column and it is fractionally distilled. The gases can be separated because they have different boiling points.
- The gases are then stored separately in large tanks and cylinders.

It should be noted that the noble gases neon, argon, krypton and xenon are obtained by this method; however, helium is more profitably obtained from natural gas.

Questions

1 Use information given in the text to construct a flow chart to show the processes involved in the extraction of gases from air.
2 Why do you think that it is more profitable to obtain gases like helium from natural gas?
3 Which gas would be the first to evaporate from liquid air, and which the last, if the lowest temperature in the process is –250 °C?

Some uses of the gases from the air

Our everyday life would not be the same without the gases we obtain from the air. Tables 11.4 and 11.5 show some of these uses.

Table 11.4 Uses of oxygen and nitrogen from the air.

Gas	Uses
Oxygen	• In steel production (Chapter 10, p. 165) • In welding to produce very hot flames • In hospitals by patients with breathing difficulties (Figure 11.9) • By mountaineers and divers • In hydrogen–oxygen fuelled rockets • By astronauts on space missions and by fire-fighters • In fuel cells • To restore life to polluted lakes and rivers
Nitrogen	• In the production of ammonia by the Haber process (see p. 176); the ammonia is then used to make nitric acid, which is used in the manufacture of dyes, explosives and fertilisers • In liquid form, as a refrigerant • As an inert atmosphere for some processes and chemical reactions, because of its unreactive nature; for example, empty oil tankers are filled with nitrogen to prevent fires • In food packaging to keep the food fresh, for example in crisp packets where it also prevents the crisps being crushed (Figure 11.10)

Table 11.5 Uses of the noble gases.

Gas	Boiling point/°C
Argon	• To fill the space between the panes of modern double glazing units • To provide an inert atmosphere in arc welding and in the production of titanium metal • To a decreasing extent, to fill incandescent light bulbs (it prevented the tungsten filament from reacting with oxygen in the air and forming the oxide)
Neon	• In advertising signs, because it glows red when electricity is passed through it • In the helium–neon gas laser (Figure 11.11) • In Geiger–Müller tubes, which are used for the detection of radioactivity
Helium	• To provide an inert atmosphere for welding • As a coolant in nuclear reactors • With 20% oxygen, as a breathing gas for deep-sea divers • To inflate the tyres of large aircraft • To fill airships and weather balloons (Figure 11.12) • In the helium–neon laser • In low-temperature research, because of its low boiling point
Krypton	• In lamps used in photographic flash units • In stroboscopic lamps • In lamps used in lighthouses

Figure 11.9 Incubators have their own oxygen supply for babies.

Figure 11.10 Inert nitrogen gas is used in food packaging.

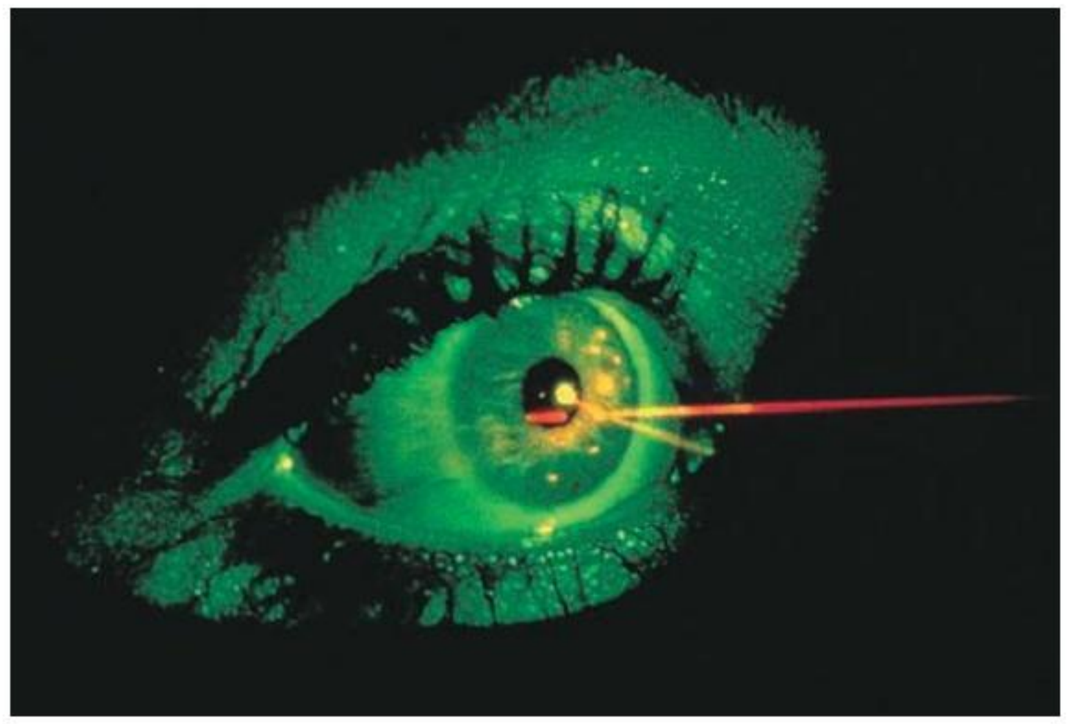


Figure 11.11 A helium–neon laser used in eye surgery.

Figure 11.12 Helium is used to fill this airship as it has a low density and is unreactive.

Questions

1 How does oxygen help to restore life to polluted lakes?
2 Why is it important to have nitrogen in fertilisers?
3 Why is helium needed to produce an inert atmosphere for welding?

Ammonia – an important nitrogen-containing chemical

Nitrogen from the air is used to manufacture ammonia, a very important **bulk chemical**. A bulk chemical is one that, because of its large usage

across a range of uses, is produced in very large quantities. The major process used for making ammonia is the **Haber process**. This process was developed by the German scientist Fritz Haber in 1913 (Figure 11.13). He was awarded a Nobel Prize in 1918 for his work. The process involves reacting nitrogen and hydrogen. It was first developed to satisfy the need for explosives during World War I, as explosives can be made from ammonia. We now have many more uses for this important gas including the manufacture of nitric acid and of fertilisers such as ammonium nitrate.

Figure 11.13 Fritz Haber (1868–1934).

Obtaining nitrogen

The nitrogen needed in the Haber process is obtained from the atmosphere by fractional distillation of liquid air (p. 174).

Obtaining hydrogen

The hydrogen needed in the Haber process is obtained from the reaction between methane and steam.

methane + steam ⇌ hydrogen + carbon monoxide

$$CH_4(g) + H_2O(g) \rightleftharpoons 3H_2(g) + CO(g)$$

This process is known as **steam re-forming**.

This reaction is a **reversible** reaction and special conditions are employed to ensure that the reaction proceeds to the right (the forward reaction), producing hydrogen and carbon monoxide. The process is carried out at a temperature of 750 °C, at a pressure of 30 atmospheres with a catalyst of nickel. These conditions enable the maximum amount of hydrogen to be produced at an economic cost.

The carbon monoxide produced is then allowed to reduce some of the unreacted steam to produce more hydrogen gas.

carbon monoxide + steam ⇌ hydrogen + carbon dioxide

$$CO(g) + H_2O(g) \rightleftharpoons H_2(g) + CO_2(g)$$

Making ammonia

In the Haber process, nitrogen and hydrogen in the correct proportions (1 : 3) are pressurised to approximately 200 atmospheres and passed over a catalyst of freshly produced, finely divided iron at a temperature of between 350 °C and 500 °C. The reaction in the Haber process is:

nitrogen + hydrogen ⇌ ammonia

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92\,\text{kJ mol}^{-1}$$

The reaction is exothermic.

The industrial conditions employed ensure that sufficient ammonia is produced at a fast enough rate. The principles which underpin equilibrium processes such as this are discussed below and in Chapter 12 (the Contact process for manufacturing sulfuric acid, p. 199).

Under these conditions the gas mixture leaving the reaction vessel contains about 15% ammonia, which is removed by cooling and condensing it as a liquid. The unreacted nitrogen and hydrogen are re-circulated into the reaction vessel to react together once more to produce further quantities of ammonia.

The 15% of ammonia produced does not seem a great deal. The reason for this is the reversible nature of the reaction. Once the ammonia is made from nitrogen and hydrogen, it decomposes to produce nitrogen and hydrogen. There comes a point when the rate at which the nitrogen and hydrogen react to produce ammonia is equal to the rate at which the ammonia decomposes. This situation is called a **chemical equilibrium**. Because the processes

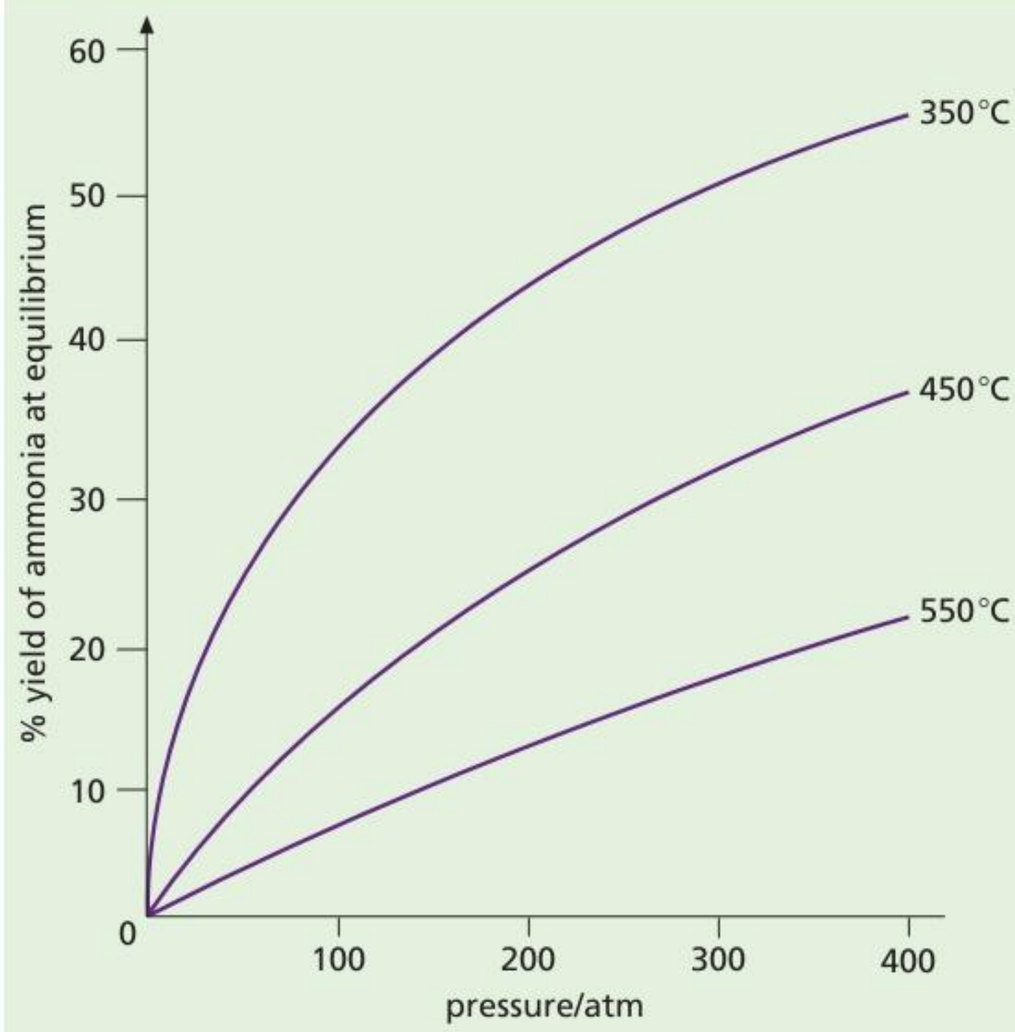


Figure 11.14 Yields from the Haber process.

continue to happen, the equilibrium is said to be **dynamic**. The conditions used ensure that the ammonia is made economically. Figure 11.14 shows how the percentage of ammonia produced varies with the use of different temperatures and pressures.

You will notice that the higher the pressure and the lower the temperature used, the more ammonia is produced. Relationships such as this were initially observed by Henri Le Chatelier, a French scientist, in 1888 (Figure 11.15). He noticed that if the pressure was increased in reactions involving gases, the reaction which produced the fewest molecules of gas was favoured. If you look at the reaction for the Haber process you will see that, going from left to right, the number of molecules of gas goes from four to two. This is why the Haber process is carried out at high pressures. Le Chatelier also noticed that reactions which were exothermic produced more products if the temperature was low. Indeed, if the Haber process is carried out at room temperature you get a higher percentage of ammonia. However, in practice the rate of the reaction is lowered too much and the ammonia is not produced quickly enough for the process to be economical. An **optimum temperature** is used to produce enough ammonia at an acceptable rate. It should be noted, however, that the increased pressure used is very expensive in capital terms and so alternative, less expensive routes involving biotechnology are being sought at the present time.

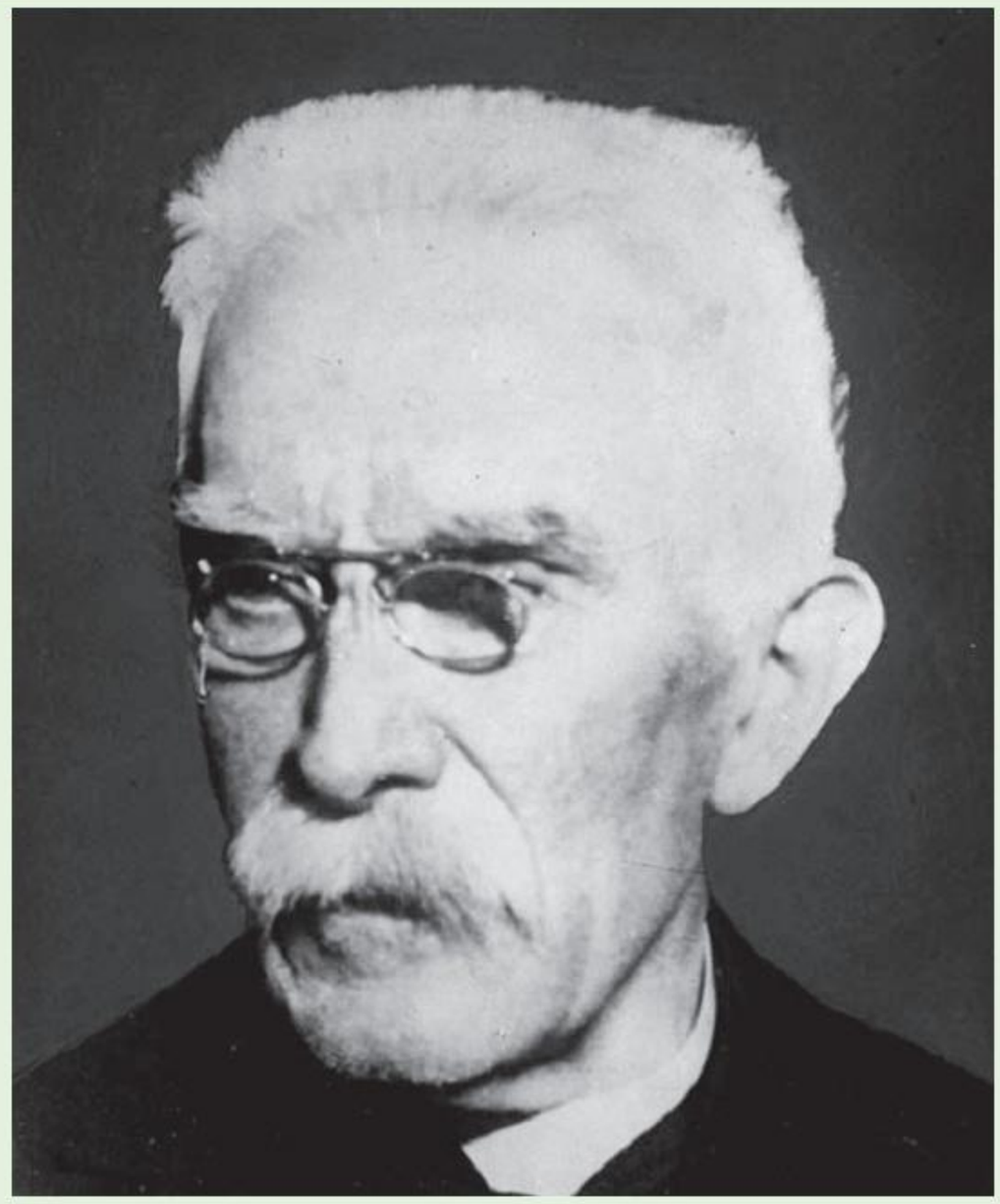


Figure 11.15 Henri Le Chatelier (1850–1936).

Worldwide, in excess of 140 million tonnes of ammonia are produced by the Haber process each year.

Questions

1 What problems do the builders of a chemical plant to produce ammonia have to consider when they start to build such a plant?

2 What problems are associated with building a plant which uses such high pressures as those required in the Haber process?

Making ammonia in the laboratory

Small quantities of ammonia gas can be produced by heating any ammonium salt, such as ammonium chloride, with an alkali, such as calcium hydroxide.

calcium hydroxide + ammonium chloride → calcium chloride + water + ammonia

$$Ca(OH)_2(s) + 2NH_4Cl(s) \rightarrow CaCl_2(s) + 2H_2O(g) + 2NH_3(g)$$

Water vapour is removed from the ammonia gas by passing the gas formed through a drying tower containing calcium oxide (Figure 11.16).

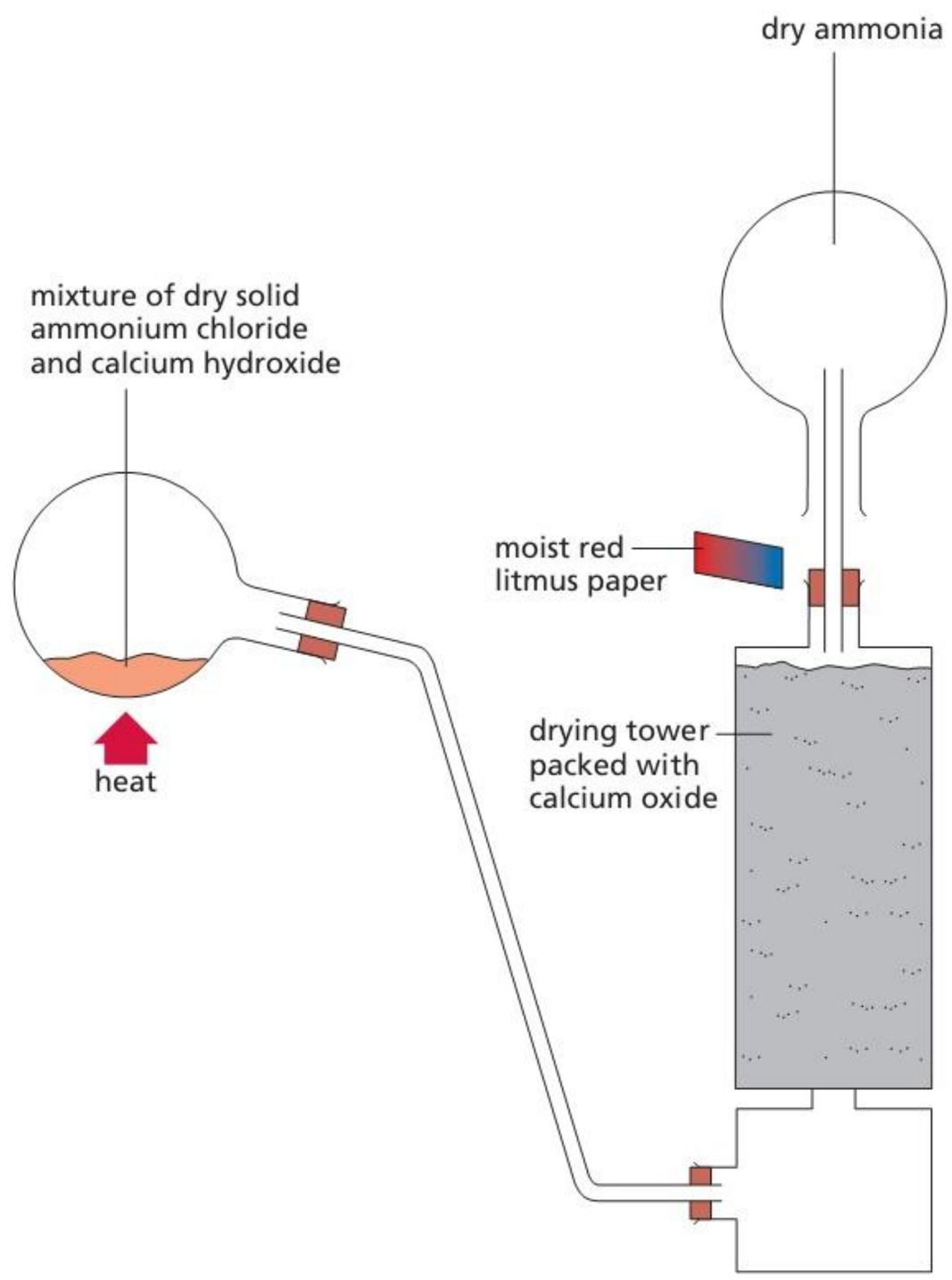


Figure 11.16 Laboratory production of ammonia gas.

This reaction forms the basis of a chemical test to show that a compound contains the ammonium ion (NH_4^+). If any compound containing the ammonium ion is heated with sodium hydroxide, ammonia gas is given off which turns damp red litmus paper blue.

Physical properties of ammonia

Ammonia (Figure 11.17):

- is a colourless gas
- is less dense than air
- has a sharp or pungent smell
- is very soluble in water with about 680 cm^3 of ammonia in each 1 cm^3 of water (at 20 °C).

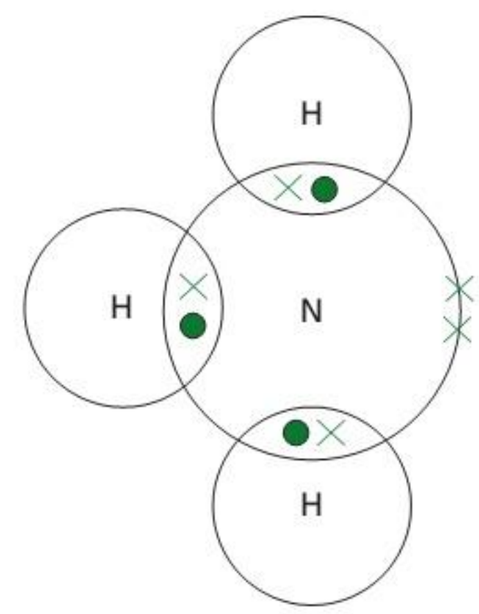


Figure 11.17 The ammonia molecule.

Chemical properties of ammonia

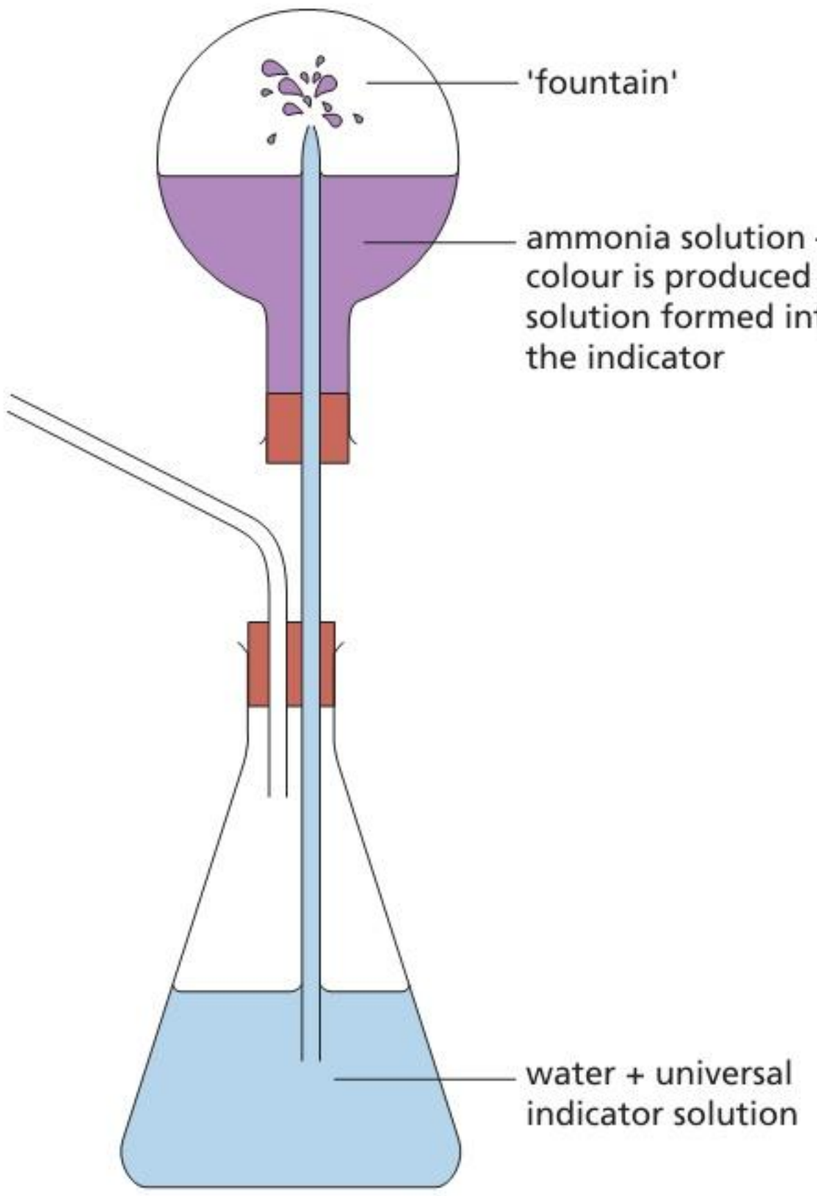
The reason ammonia is so soluble in water is that some of it reacts with the water. The high solubility can be shown by the 'fountain flask experiment' (Figure 11.18, p. 180). As the first drop of water reaches the top of the tube all the ammonia gas in the flask dissolves, creating a much reduced pressure. Water then rushes up the tube to fill the space once occupied by the dissolved gas. This creates the fountain.

If the water initially contained some universal indicator, then you would also see a change from green to blue when it comes into contact with the dissolved ammonia. This shows that ammonia solution is a weak alkali, although dry ammonia gas is not. This is because a little of the ammonia gas has reacted with the water, producing ammonium ions and hydroxide ions. The hydroxide ions produced make the solution of ammonia alkaline.

ammonia	+	water	⇌	ammonium ions	+	hydroxide ions
$NH_3(g)$	+	$H_2O(l)$	⇌	$NH_4^+(aq)$	+	$OH^-(aq)$

The solution is only weakly alkaline because of the reversible nature of this reaction, which results in a relatively low concentration of hydroxide ions. Ammonia gas dissolved in water is usually known as aqueous ammonia.

Aqueous ammonia can be used to identify salts of Cu^{2+}, Fe^{2+}, Fe^{3+}, Al^{3+}, Zn^{2+}, Cr^{3+} and Ca^{2+} ions. The colour of the precipitate or solution formed identifies the metal present (Table 11.6).



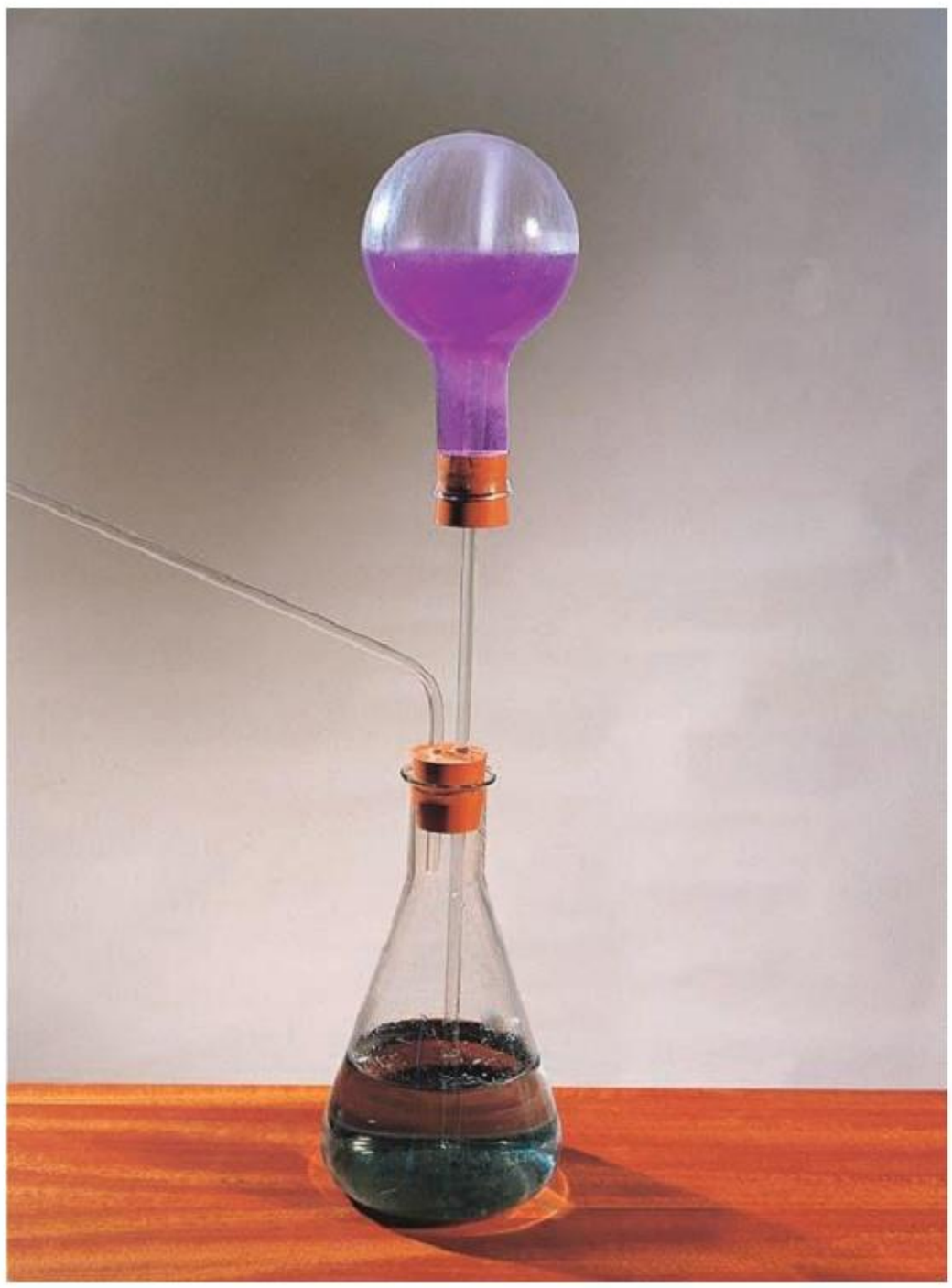


Figure 11.18 The fountain flask experiment.

Table 11.6 Identifying metal ions using aqueous ammonia.

Metal ion	With a few drops of ammonia solution	With excess ammonia solution
$Cu^{2+}(aq)$	Gelatinous blue precipitate	Precipitate dissolves to give a deep blue solution
$Fe^{2+}(aq)$	Dirty green precipitate	Dirty green precipitate remains
$Fe^{3+}(aq)$	Rust brown precipitate	Rust brown precipitate remains
$Al^{3+}(aq)$	White precipitate	White precipitate remains
$Zn^{2+}(aq)$	White precipitate	White precipitate dissolves to give a colourless solution
$Ca^{2+}(aq)$	Faint white precipitate	Faint white precipitate remains
$Cr^{3+}(aq)$	Light green precipitate	Precipitate dissolves to some extent

Questions

1. Calcium oxide is used to dry ammonia gas in its laboratory preparation. Write a word and balanced chemical equation to show how calcium oxide can react with the water vapour to remove it from damp ammonia gas.
2. Explain why ammonia gas only acts as a weak alkali in the presence of water.
3. Describe how you could use aqueous ammonia to distinguish between three unlabelled bottles containing solutions of $Al^{3+}(aq)$, $Ca^{2+}(aq)$ and $Cu^{2+}(aq)$.

Artificial fertilisers

Some of the ammonia produced by the Haber process is used to produce nitric acid. If ammonia is then reacted with the nitric acid, we have the basic reaction for the production of many artificial fertilisers. The use of artificial fertilisers is essential if farmers are to produce sufficient crops to feed the ever-increasing world population. Crops remove nutrients from the soil as they grow; these include nitrogen, phosphorus and potassium. Artificial fertilisers are added to the soil to replace these nutrients and others, such as calcium, magnesium, sodium, sulfur, copper and iron. Examples of nitrogenous fertilisers (those which contain nitrogen) are shown in Table 11.7.

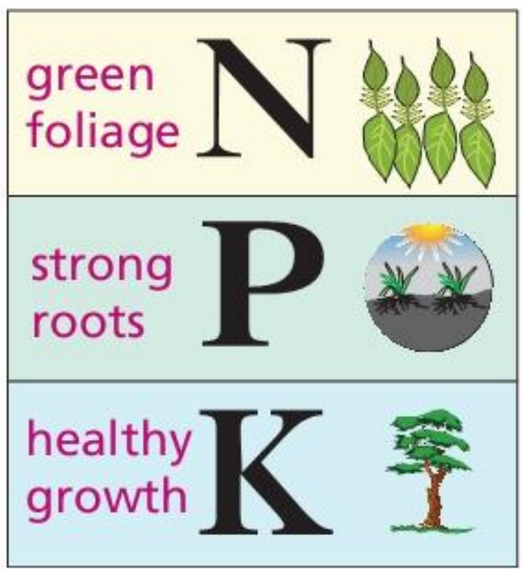
Table 11.7 Some nitrogenous fertilisers.

Fertiliser	Formula
Ammonium nitrate	NH_4NO_3
Ammonium phosphate	$(NH_4)_3PO_4$
Ammonium sulfate	$(NH_4)_2SO_4$
Urea	$CO(NH_2)_2$

Artificial fertilisers can also make fertile land which was once unable to support crop growth. The fertilisers which add the three main nutrients (N, P and K) are called NPK fertilisers. They contain ammonium nitrate

(NH_4NO_3), ammonium phosphate ($(NH_4)_3PO_4$) and potassium chloride (KCl) in varying proportions (Figure 11.19). Fertilisers have an important role in the nitrogen cycle (see below).

a Different fertilisers contain differing amounts of the elements nitrogen, phosphorus and potassium.



b The different NPK elements are responsible for the healthy growth of plants in different ways.

Figure 11.19

Figure 11.20 Fertilisers have been used to help create some of the best fruit on sale.

Manufacture of ammonium nitrate

Ammonium nitrate (Nitram®) is probably the most widely used nitrogenous fertiliser. It is manufactured by reacting ammonia gas and nitric acid.

$$\text{ammonia} + \text{nitric acid} \rightarrow \text{ammonium nitrate}$$
$$NH_3(g) + HNO_3(aq) \rightarrow NH_4NO_3(aq)$$

The nitrogen cycle

The vital importance of nitrogen to both plants and animals can be summarised by the **nitrogen cycle** (Figure 11.21).

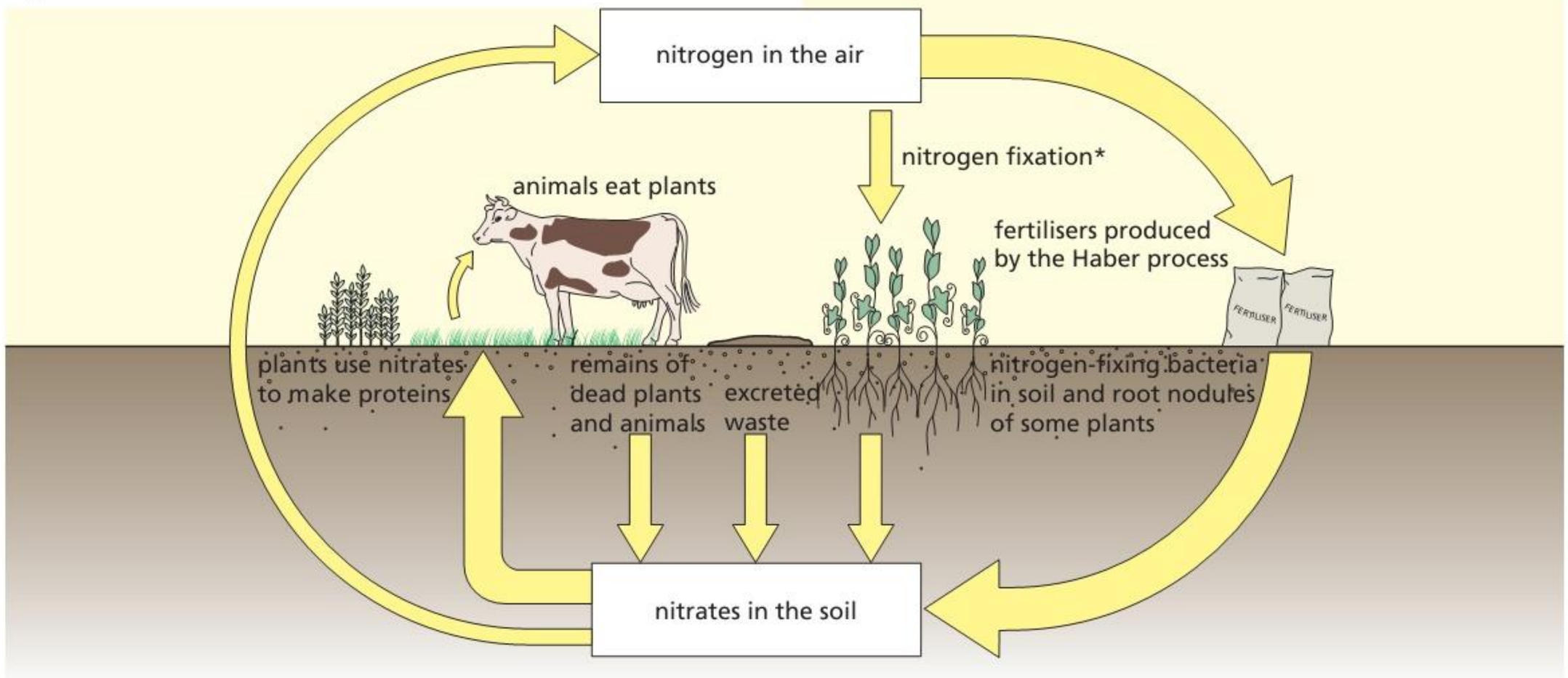


Figure 11.21 The nitrogen cycle.
***Nitrogen fixation** is the direct use of atmospheric nitrogen in the formation of important compounds of nitrogen. Bacteria present in the root nodules of certain plants are able to take nitrogen directly from the atmosphere to form essential protein molecules.

If farm crops are harvested from the land rather than left to decay, the soil becomes deficient in this important element. The nitrogen is removed in the harvested crops rather than remaining as the plants decay. In addition, nitrates can be washed from the soil by the action of rain (leaching). For the soil to remain fertile for the next crop, the nitrates need to be replaced. The natural process is by decay or by the action of lightning on atmospheric nitrogen. Without the decay, however, the latter process is not efficient enough to produce nitrates on the scale required.

Farmers often need to add substances containing these nitrates. Such substances include farmyard manure and artificial fertilisers. One of the most commonly used artificial fertilisers is **ammonium nitrate**, which as you saw earlier is made from ammonia gas and nitric acid, both nitrogen-containing compounds.

Problems with fertilisers

If artificial fertilisers of all kinds are not used correctly, problems can arise. If too much fertiliser is applied to the land, rain washes the fertiliser off the land and into rivers and streams. This leaching leads to **eutrophication** by encouraging the growth of algae and marine plants. As the algae die and decay, oxygen is removed from the water, leaving insufficient amounts for fish and other organisms to survive (Figure 11.22). In extreme cases, no normal aquatic life can survive. There are also worries about the effect of agricultural fertilisers, especially nitrates, on the public water supply. There is evidence that the fertiliser sodium nitrate is implicated in infant brain damage, and it is also suspected of being carcinogenic.

Figure 11.22 Over-use of fertilisers has led to eutrophication in this river. Oxygen is removed from the water and organisms cannot survive.

Questions

1. Calculate the percentage of nitrogen in each of the four fertilisers in Table 11.7. (A_r: H = 1; N = 14; O = 16; P = 31; S = 32)
2. Write down a method that you could carry out in a school laboratory to prepare a sample of ammonium sulfate fertiliser.
3. Use your research skills to find out about the process of producing nitric acid by the oxidation of ammonia. Include in your answer key chemical equations and also refer to the major uses of this acid.

Atmospheric pollution

The two major resources considered in this chapter, water and air, are essential to our way of life and our very existence. Water and air make up the environment of a living organism. The environment is everything in the surroundings of an organism that could possibly influence it. Humans continually pollute these resources. We now look at the effects of the various sources of pollution of the air and at the methods used to control or eliminate them. For a discussion of water pollution, see p. 190.

Air pollution is all around us. Concentrations of gases in the atmosphere such as carbon monoxide, sulfur dioxide and nitrogen oxides are increasing with the increasing population. As the population rises there is a consequent increase in the need for energy, industries and motor vehicles. These gases are produced primarily from the combustion of the fossil fuels coal, oil and gas, but they are also produced by the smoking of cigarettes.

Motor vehicles are responsible for much of the air pollution in large towns and cities. They produce four particularly harmful pollutants:

- carbon monoxide – toxic gas
- sulfur dioxide – major irritant and constituent of acid rain
- hydrocarbons – carcinogenic
- oxides of nitrogen – major irritant, acid rain and photochemical smog
- lead compounds – toxic.

Nitrogen monoxide is formed by the reaction of nitrogen and oxygen inside the internal combustion engine at high temperatures.

$$\text{nitrogen} + \text{oxygen} \rightarrow \text{nitrogen monoxide}$$
$$N_2(g) + O_2(g) \rightarrow 2NO(g)$$

The nitrogen monoxide, emitted as an exhaust gas, reacts with oxygen from the air to form the

brown acidic gas nitrogen(IV) oxide (nitrogen dioxide).

nitrogen monoxide	+	oxygen	→	nitrogen(IV) oxide
$2NO(g)$	+	$O_2(g)$	→	$2NO_2(g)$

This dissolves in water to produce nitric acid (this is a contributor to acid rain; see Chapter 12 p. 198). Many non-metal oxides, such as nitrogen dioxide and carbon dioxide (see Chapter 13 p. 214), form acidic solutions.

Concern about pollution due to cars has led to the introduction of strict regulations and now all new cars must have a device known as a catalytic converter fitted to eliminate the production of some of these gases.

The catalytic converter acts as a device to speed up reactions which involve the pollutant gases, converting them to less harmful products, such as nitrogen and carbon dioxide. It should be noted that catalytic converters can only be used with unleaded petrol as the lead 'poisons' the catalyst, preventing it from catalysing the reactions. For a further discussion of catalytic converters see Chapter 7, p. 110.

Another method that has been introduced to reduce the amount of pollutants is that of the 'lean burn' engine. Although this type of engine reduces the amounts of carbon monoxide and oxides of nitrogen produced, it actually increases the amount of hydrocarbons in the exhaust gases.

A further method of regulating pollutant gases is to convert petrol-burning engines to LPG (liquid petroleum gas) engines, whilst retaining the ability to burn petrol. These cars are known as dual-fuel cars. Much research is directed towards an efficient electric motor that can be fitted in place of the petrol engine in a car. This development is moving forward at a pace.

Internationally, the use of lead compounds as additives remains an issue because of the continued use of leaded petrol. However, the concern about these additives is lessening as worldwide recognition of the problems of lead pollution in the environment and its effects on living systems is leading to their being removed from petrol. Perhaps of greater concern globally will be the increase of **particulates** in the air (Figure 11.23). ('Particulates' is a general term used to describe very small particles in the atmosphere, such as certain types of smoke emitted from diesel engines, as well as dust. These particulates have been associated with a variety of lung complaints in both adults and children.)

Figure 11.23 Particulates produced by diesel engines are a real health problem worldwide.

Heavy industry (Figure 11.24) and power stations are major sources of sulfur dioxide, formed by the combustion of coal, oil and gas, which contain small amounts of sulfur.

sulfur	+	oxygen	→	sulfur dioxide
$S(s)$	+	$O_2(g)$	→	$SO_2(g)$

Figure 11.24 Sulfur dioxide is a major pollutant produced by industry.

This sulfur dioxide gas dissolves in rainwater to form the weak acid, sulfurous acid (H_2SO_3).

sulfur dioxide	+	water	⇌	sulfurous acid
$SO_2(g)$	+	$H_2O(l)$	⇌	$H_2SO_3(aq)$

A further reaction occurs in which the sulfurous acid is oxidised to sulfuric acid. Solutions of these acids are the principal contributors to acid rain. For a further discussion of acid rain, see Chapter 12, p. 198.

Units called flue gas desulfurisation (FGD) units are being fitted to some power stations throughout the world to prevent the emission of sulfur dioxide gas. Here, the sulfur dioxide gas is removed from the waste gases by passing them through calcium hydroxide slurry. This not only removes the sulfur dioxide but also creates calcium sulfate, which can be sold to produce plasterboard (Figure 11.25). The FGD units are very expensive and therefore the sale of the calcium sulfate is an important economic part of the process.

Figure 11.25 This plasterboard is made using calcium sulfate from an FGD plant.

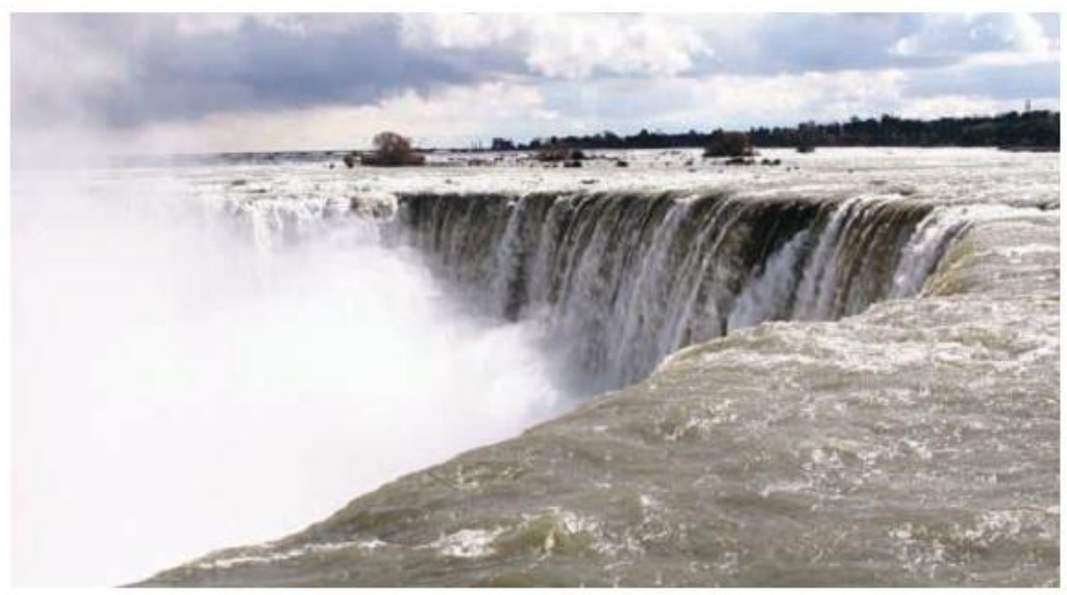
Another source of pollution is hydrogen sulfide (H_2S), a poisonous gas which has the distinctive smell of rotten eggs. It is present in volcanic emissions. Crude oil also contains small quantities of hydrogen sulfide and this is removed during its processing. Since the introduction of catalytic converters, modern cars have shown an increased tendency to produce H_2S on heavy acceleration. This can be improved by tuning the engine. Where amounts of hydrogen sulfide are present in the air, it causes irritation to the lungs. It also blackens lead-based paints.

Questions

1. Write a balanced chemical equation to represent the reaction which takes place between sulfur dioxide and calcium hydroxide slurry in the FGD unit of a power station.
2. It has been found that air pollution in cities damages lung development in children. Use the information in the text and any other information to make a list of the pollutants responsible for this problem.
3. Write down one problem that can be caused by each of these air pollutants:
   a nitrogen dioxide
   b particulates.
4. Is carbon dioxide an atmospheric pollutant? Give reasons for your answer.

Water

Water is the commonest compound on this planet. More than 70% of the Earth's surface is covered with sea, and the land masses are dotted with rivers and lakes (Figure 11.26a). It is vital to our existence and survival because it is one of the main constituents in all living organisms. For example, your bones contain 72% water, your kidneys are about 82% water and your blood is about 90% water (Figure 11.26b).



a Millions of tonnes of water pass over this waterfall every day.

b Your blood contains a lot of water.
Figure 11.26

Water has many other important uses besides sustaining life. These include:

In the home:

- cooking
- cleaning
- drinking.

In industry:

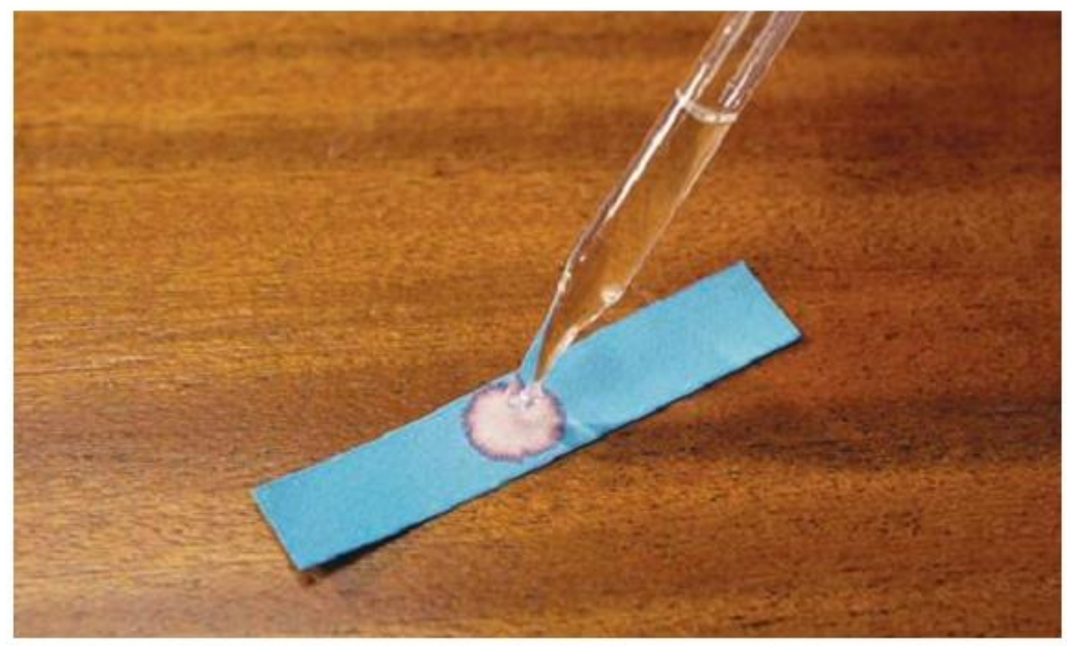
- as a solvent
- as a coolant
- for cleaning
- as a chemical reactant.

Water is a neutral, colourless liquid which (at 1 atmosphere pressure) boils at 100 °C and freezes at 0 °C (Figure 11.27).

Figure 11.27 Liquid water turns to steam at 100 °C and to ice at 0 °C.

You can find out whether a colourless liquid contains water by adding the unknown liquid to anhydrous copper(II) sulfate. If this changes from white to blue, then the liquid contains water (Figure 11.28a).

a Anhydrous copper(II) sulfate goes blue when water is added to it.



b Cobalt chloride paper turns pink when water is dropped on to it.
Figure 11.28 Tests for the presence of water.

Another test is to dip blue cobalt chloride paper into the liquid. If the paper turns pink, then the liquid contains water (Figure 11.28b).

You have already seen in Chapter 5 that water may be electrolysed (when acidified with a little dilute sulfuric acid). When this is done, the ratio of the volume of the gas produced at the cathode to that produced at the anode is 2 : 1. This is what you might expect, since the formula of water is H_2O!

The unique properties of water

Water is a unique substance. Not only is it an excellent solvent for many ionic substances, such as sodium chloride, but it also has some unusual properties. For example:

- It has an unusually high boiling point for a molecule of its relatively low molecular mass.
- It has a greater specific heat capacity than almost any other liquid.
- It decreases in density when it freezes (Figure 11.29).

Figure 11.29 When water freezes, its density falls; this is why icebergs float.

These unusual properties can be explained by **hydrogen bonding**. This is a weak intermolecular force (bond) which occurs between water molecules because the bonds within the molecules are polar. A polar bond is one which contains, for example, oxygen attached to hydrogen by a covalent bond. The shared electrons in the bond are drawn towards the oxygen atom to create a small negative charge on the oxygen atom (δ–). There is then an equally small positive charge on the hydrogen atom (δ+). The water molecules then attract one another as shown in Figure 11.30. In the case of water, this attraction is called a hydrogen bond. It is a much weaker bond than a covalent bond.

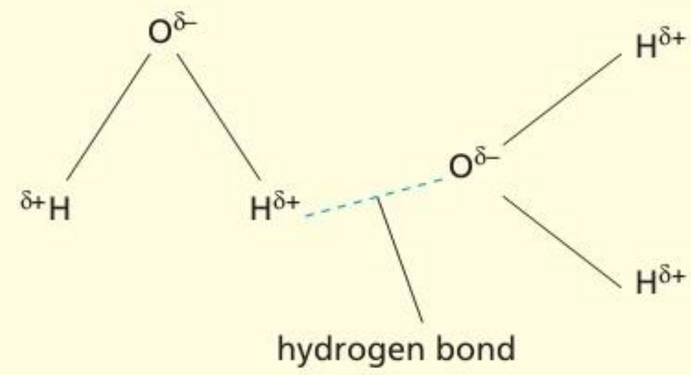


Figure 11.30 The unusual properties of water can be explained by hydrogen bonding.

The fact that water is such an effective solvent can cause some problems. These include the formation of hard water and leaching of fertilisers (Figure 11.22, p. 182). Hard water can cause further problems and must in certain circumstances undergo a softening process. This is discussed in detail on p. 188.

The water cycle

Water circulates around the Earth. The way it does this can be described by the **water cycle** (Figure 11.31). The driving force for the water cycle is the Sun.

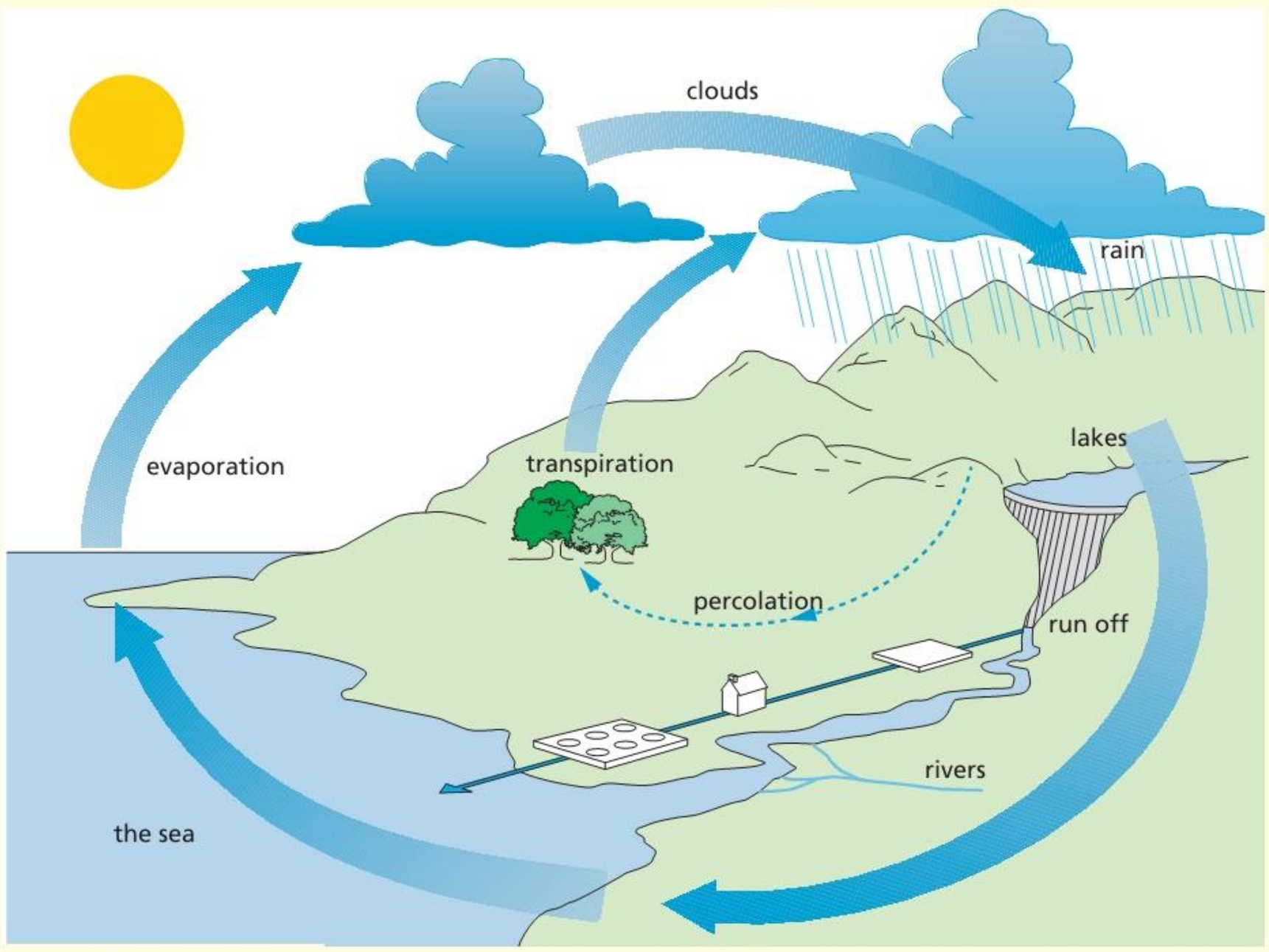


Figure 11.31 The water cycle.

- Heat from the Sun causes evaporation from oceans, seas and lakes. Water vapour is also formed from the evaporation of water from leaves (transpiration), through respiration and through combustion. The water vapour rises and cools, and condenses to form tiny droplets of water. These droplets form clouds.
- The clouds are moved along by air currents. As they cool, the tiny droplets join to form larger droplets, which fall as rain when they reach a certain size.
- The water that falls as rain runs into streams and rivers and then on into lakes, seas and oceans.

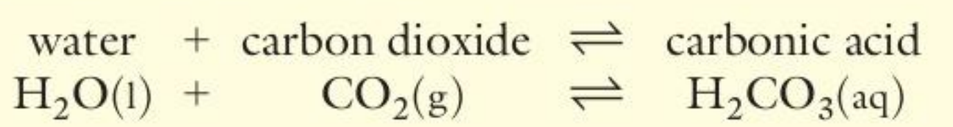
Some water is diverted from rivers into reservoirs and purified (p. 190). This water is then used in industry and the home. After use it is dirty and has to be treated at the sewage works before being returned to a river or the sea. Eventually, all the water finds its way into the sea. In this way all the water that left the sea eventually returns there, and the cycle can continue.

Question

1 Construct a simplified version of the water cycle using 'key words' in boxes and the 'processes involved' over linking arrows.

Hardness in water

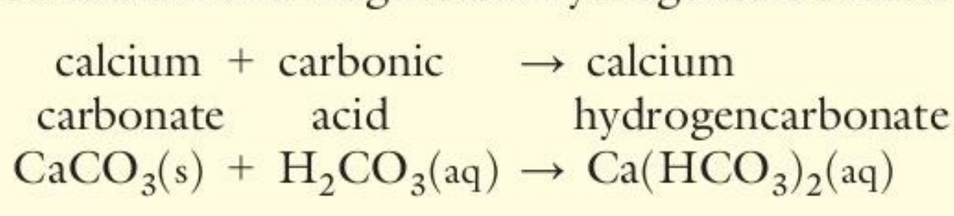
The water you drink contains some dissolved solids and gases. This dissolved material usually is not harmful and can, in fact, be good for you. Where do these solids come from? Rainwater dissolves carbon dioxide as it falls through the atmosphere. A small fraction of this dissolved carbon dioxide reacts with the water to produce carbonic acid, which is a weak acid (see Chapter 8, p. 120).



water + carbon dioxide ⇌ carbonic acid

$H_2O(l) + CO_2(g) \rightleftharpoons H_2CO_3(aq)$

As this solution passes over and through rocks containing limestone (calcium carbonate, $CaCO_3$) and dolomite (magnesium carbonate, $MgCO_3$), the weak acid in the rain attacks these rocks and very slowly dissolves them. The dissolved substances are called calcium and magnesium hydrogencarbonates.



calcium carbonate + carbonic acid → calcium hydrogencarbonate

$CaCO_3(s) + H_2CO_3(aq) \rightarrow Ca(HCO_3)_2(aq)$

Some of the rock strata may contain gypsum (calcium sulfate, $CaSO_4.2H_2O$), anhydrite ($CaSO_4$) or kieserite ($MgSO_4.H_2O$), which are very sparingly soluble in water. The presence of any of these dissolved sulfates or hydrogencarbonates causes the water to become 'hard'.

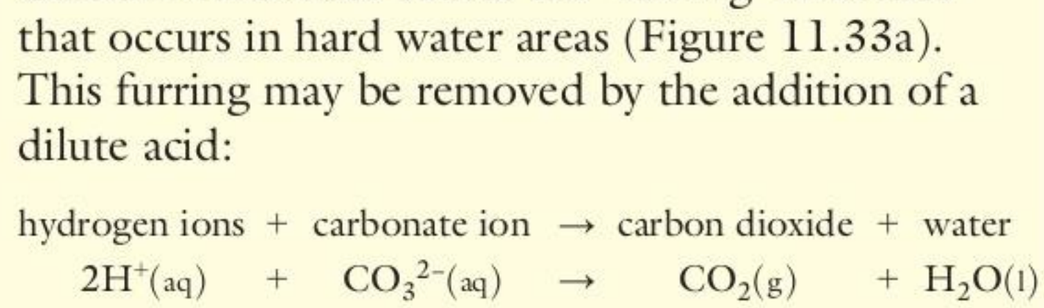
Hardness in water can be divided into two types – **temporary** and **permanent**. Temporary hardness is caused by the presence of dissolved calcium or magnesium hydrogencarbonates. Temporary hardness is so called because it is easily removed by boiling.

Permanent hardness is caused by the presence of dissolved calcium or magnesium sulfates. Permanent hardness is much more difficult to remove and certainly cannot be removed by boiling.

When water containing any of these substances is evaporated, a white solid deposit of calcium or magnesium sulfate and/or calcium carbonate is left behind (Figure 11.32).

Figure 11.32 Evaporating pure water leaves no deposit, while temporary or permanent hard water leaves white deposits behind.

Calcium carbonate causes the 'furring' in kettles that occurs in hard water areas (Figure 11.33a). This furring may be removed by the addition of a dilute acid:



hydrogen ions + carbonate ion → carbon dioxide + water

$2H^+(aq) + CO_3^{2-}(aq) \rightarrow CO_2(g) + H_2O(l)$

Blockages in hot water pipes (Figure 11.33b) are caused by a similar process to the furring of kettles. A thick deposit of limescale builds up.

a The deposit in kettles ('furring') is caused by calcium carbonate from hard water.

b The thick deposits of limescale can be seen in this hot water pipe.
Figure 11.33

Stalactites and stalagmites are found in underground caverns in limestone areas. They are formed from the slow decomposition of calcium or magnesium hydrogencarbonates in water (Figure 11.34).

calcium hydrogencarbonate	→	calcium carbonate	+	carbon dioxide	+ water
$Ca(HCO_3)_2(aq)$	→	$CaCO_3(s)$	+	$CO_2(g)$	+ $H_2O(l)$

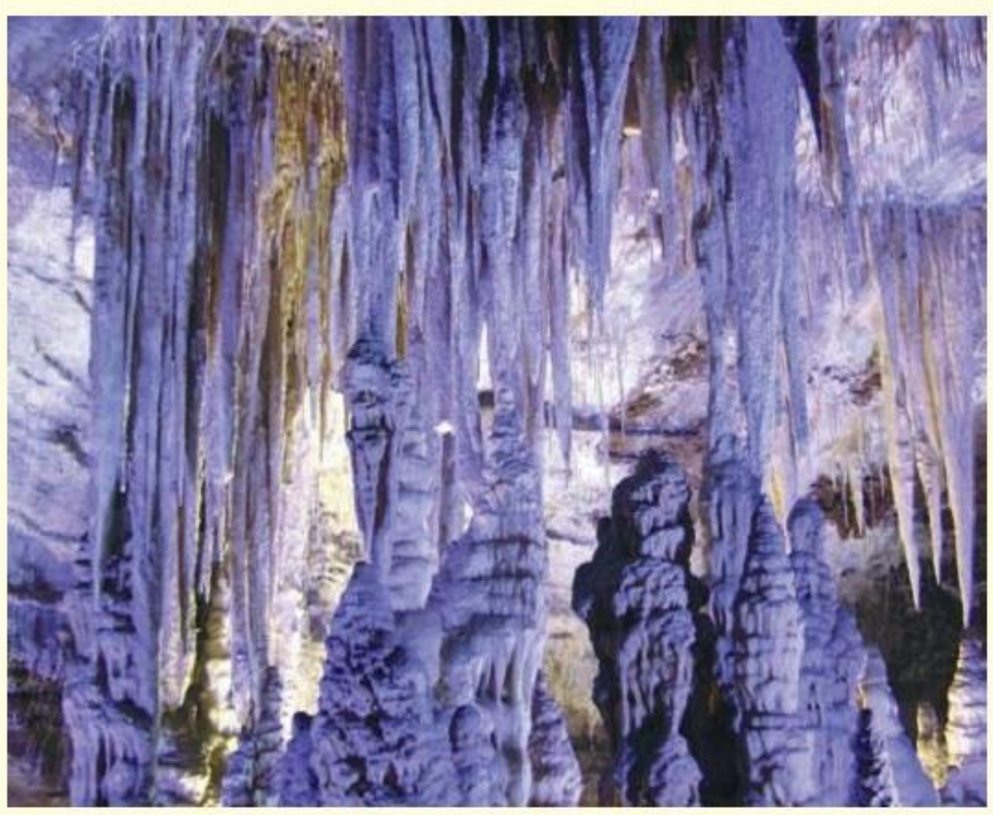


Figure 11.34 Stalactites and stalagmites have formed over hundreds of thousands of years.

Effect of hard water on soap

In hard water areas all over the world, it is difficult to make the soap lather. Instead, the water becomes cloudy. This cloudiness is caused by the presence of a solid material (a precipitate) formed by the reaction of the dissolved substances in the water with soap (basically sodium stearate) and it is a real problem. This white precipitate is known as scum (Figure 11.35).

sodium stearate (soap)	+	calcium hydrogen-carbonate	→	calcium stearate (scum)	+	sodium hydrogen-carbonate
$2NaSt(aq)$	+	$Ca(HCO_3)_2(aq)$	→	$Ca(St)_2(s)$	+	$2NaHCO_3(aq)$

St = stearate NaSt = $C_{17}H_{35}COO^-Na^+$

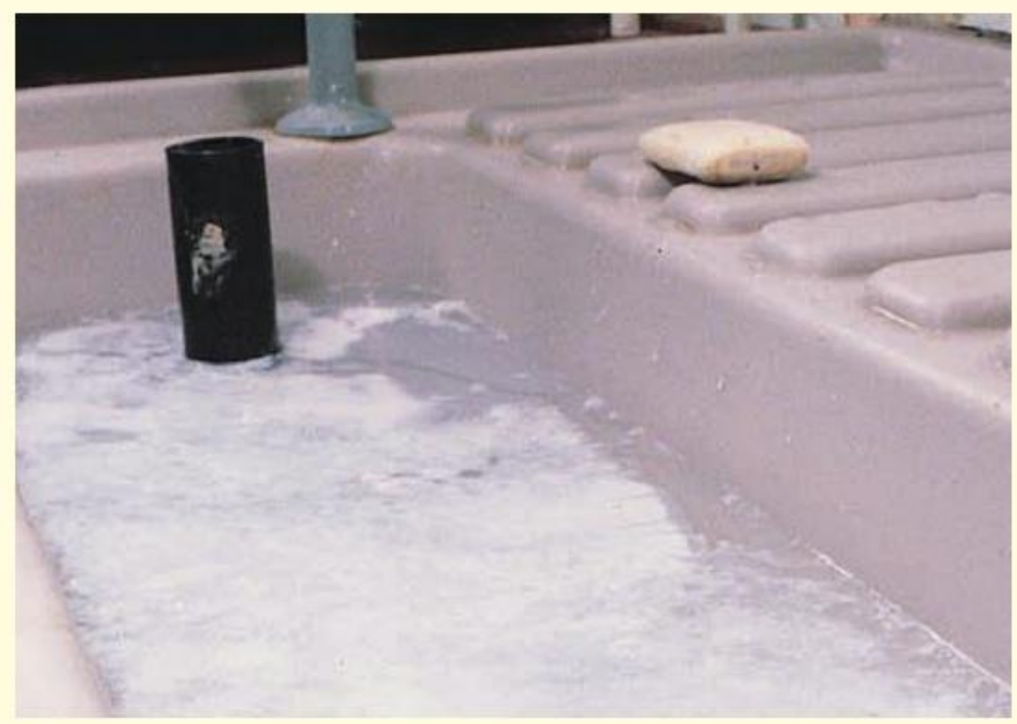


Figure 11.35 Soap and hard water form scum.

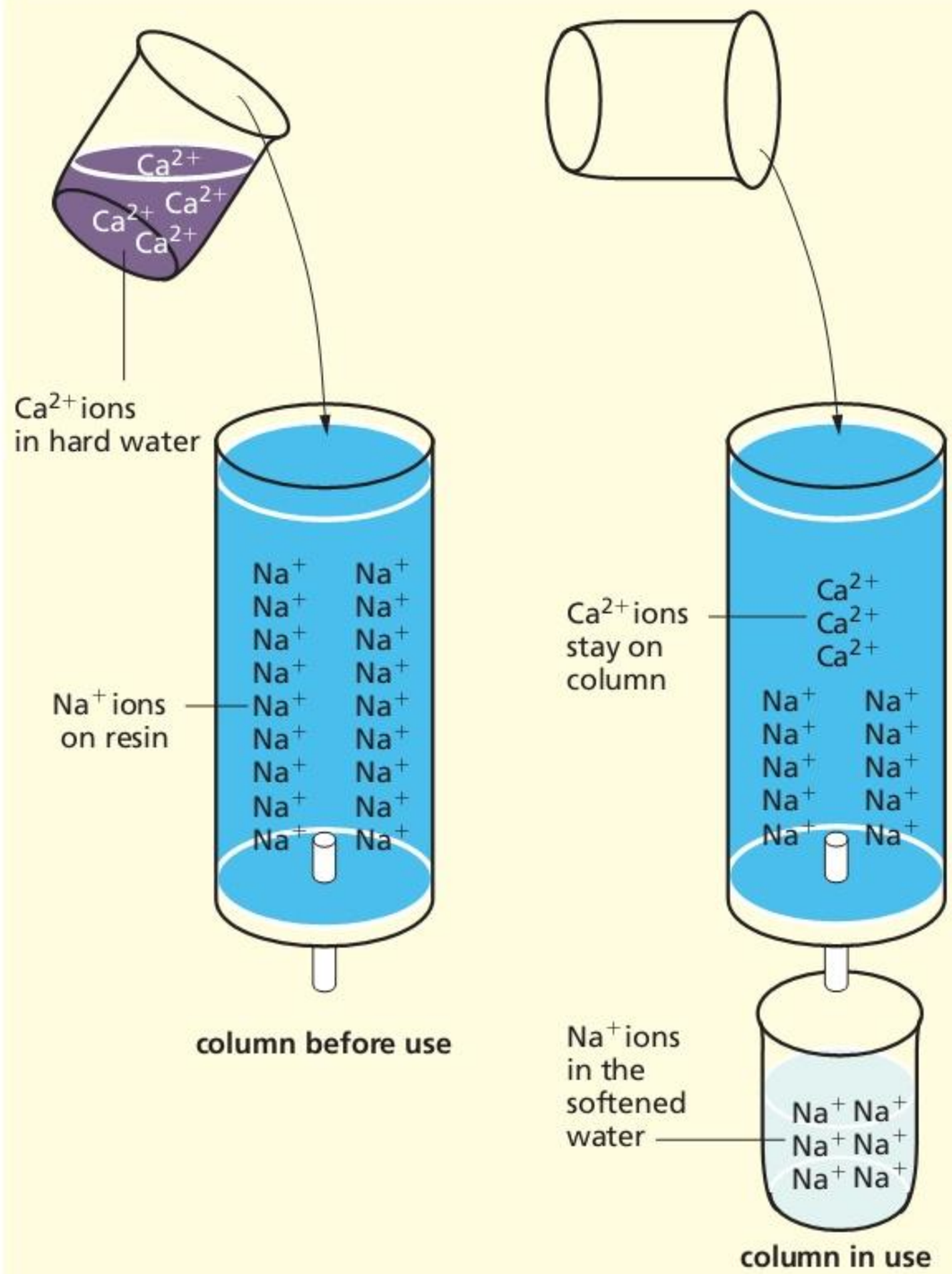
The amount of soap required to just produce a lather with the hard water can be used to estimate the hardness in water.

To overcome the problem of scum formation, soapless detergents have been developed. They do not produce a scum because they do not react with the substances in hard water. For further discussion of soapless detergents see Chapter 15, p. 239.

Removal of hardness

Temporary hardness is easily removed from water by boiling. When heated, the calcium hydrogencarbonate decomposes, producing insoluble calcium carbonate.

calcium hydrogencarbonate	$\xrightarrow{heat}$	calcium carbonate	+	water	+	carbon dioxide
$Ca(HCO_3)_2(aq)$	→	$CaCO_3(s)$	+	$H_2O(l)$	+	$CO_2(g)$



a The ion exchange process.
Figure 11.36

b Ion exchange resin.

The substances in permanently hard water are not decomposed when heated and therefore cannot be removed by boiling. Both types of hardness can be removed by any of the following methods.

- **Addition of washing soda ($Na_2CO_3.10H_2O$) crystals.** In each type, the calcium or magnesium ion, which actually causes the hardness, is removed as a precipitate and can, therefore, no longer cause hardness.

calcium ions (from hard water) + carbonate ions (from washing soda) → calcium carbonate

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$$

- **Ion exchange.** The water is passed through a container filled with a suitable resin containing sodium ions. The calcium or magnesium ions causing the hardness are exchanged for the sodium ions in the resin (Figure 11.36).

calcium ion + sodium–resin → calcium–resin + sodium ion

$$Ca^{2+}(aq) + Na_2\text{–}R(s) \rightarrow Ca\text{–}R(s) + 2Na^+(aq)$$

When all the sodium ions have been removed from the resin, it can be regenerated by pouring a solution of a suitable sodium salt through it.

- **Distillation.** The water is distilled away from the dissolved substances. This method, however, is far too expensive to be used on a large scale.

Advantages and disadvantages of hard water

Problems and advantages associated with hard water are detailed in Table 11.8.

Table 11.8 Disadvantages and advantages of hard water.

Disadvantages	Advantages
Wastes soap	Has a nice taste
Causes kettles to fur	Calcium ions in hard water are required by the body for bones and teeth
Can cause hot water pipes to block	Coats lead pipes with a thin layer of lead(II) sulfate or lead(II) carbonate and cuts down the possibility of lead poisoning
Can spoil the finish of some fabrics	Some industries favour hard water. For example, the leather industry prefers to cure leather in hard water

Questions

1 Make a list of the main methods of softening hard water. In each case write a suitable equation(s) to summarise the chemical reactions involved.

2 One of the substances found in some temporary hard waters is magnesium hydrogencarbonate. Write word and balanced chemical equations to show the effect of heat on this substance in aqueous solution.

Water pollution and treatment

An adequate supply of water is essential to the health and well-being of the world's population. Across the planet, biological and chemical pollutants are affecting the quality of our water. An adequate supply of fresh drinking water is needed for everyone on the planet. Lack of availability of fresh water leads to waterborne diseases, such as cholera and typhoid, and to diarrhoea, which is one of the biggest killers across the world.

Agriculture needs a water supply in order to irrigate crops, especially in areas of the world with hot climates. The production of more and more crops for the ever-increasing population is essential.

Water is very good at dissolving substances. Thus, it is very unusual to find really pure water on this planet. As water falls through the atmosphere, on to and then through the surface of the Earth, it dissolves a tremendous variety of substances. Chemical fertilisers washed off surrounding land will add nitrate ions (NO_3^-) and phosphate ions (PO_4^{3-}) to the water, owing to the use of artificial fertilisers such as ammonium nitrate and ammonium phosphate.

The nitrates encourage the growth of algae which eventually die and decay, removing oxygen from the water. It may also contain human waste as well as insoluble impurities such as grit and bacteria, and oil and lead 'dust' (to a decreasing extent) from the exhaust fumes of lorries and cars (Figure 11.37).

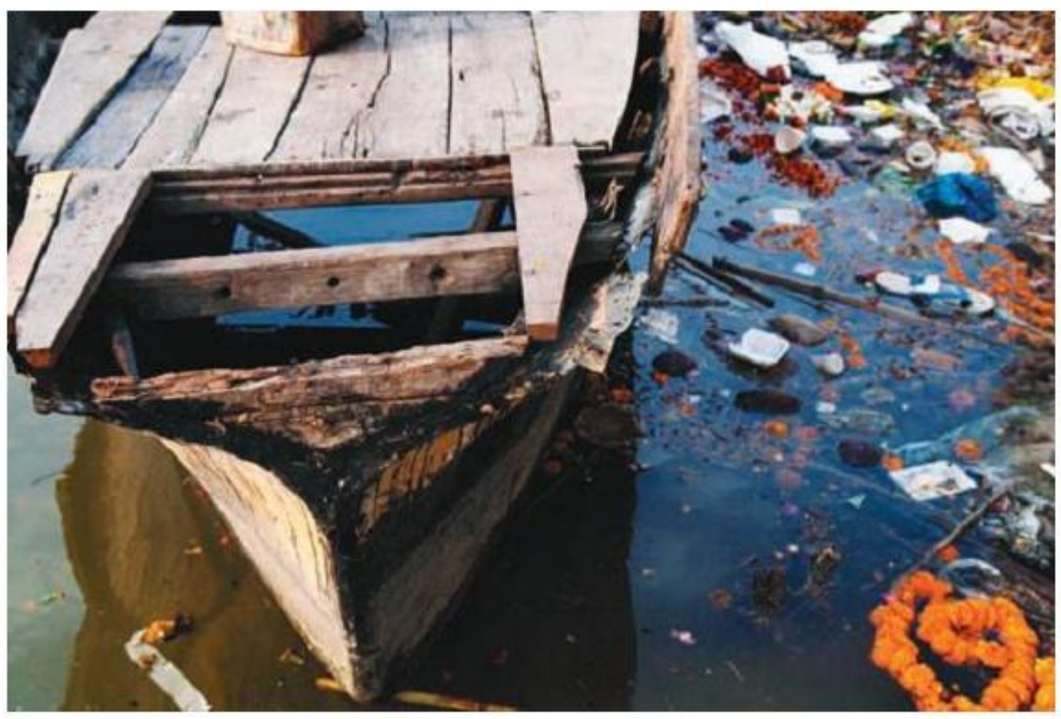


Figure 11.37 A badly polluted river.

All these artificial, as well as natural, impurities must be removed from the water before it can be used. Recent regulations in many countries have imposed strict guidelines on the amounts of various substances allowed in drinking water.

Figure 11.38 This lake is used as a source of drinking water.

A lot of drinking water is obtained from lakes and rivers where the pollution levels are low (Figure 11.38). Undesirable materials removed from water include:

- colloidal clay (clay particles in the water)
- bacteria
- chemicals which cause the water to be coloured and foul tasting
- acids, which are neutralised.

Making water fit to drink

The treatment needed to make water fit to drink depends on the source of the water. Some sources, for example mountain streams, may be almost pure and boiling may be enough to kill any micro-organisms present. However, others, such as slow-flowing rivers, may be contaminated. The object of treating contaminated water is to remove all micro-organisms that may cause disease.

The process of water treatment involves both filtration and chlorination and is summarised in Figure 11.39.

1. Impure water is passed through screens to filter out floating debris.
2. Aluminium sulfate is added to coagulate small particles of clay so that they form larger clumps, which settle more rapidly.
3. Filtration through coarse sand traps larger, insoluble particles. The sand also contains specially grown microbes which remove some of the bacteria.
4. A sedimentation tank has chemicals known as flocculants, for example aluminium sulfate, added to it to make the smaller particles (which remain in the water as colloidal clay) stick together and sink to the bottom of the tank.
5. These particles are removed by further filtration through fine sand. Sometimes a carbon slurry is used to remove unwanted tastes and odours, and a lime slurry is used to adjust the acidity.

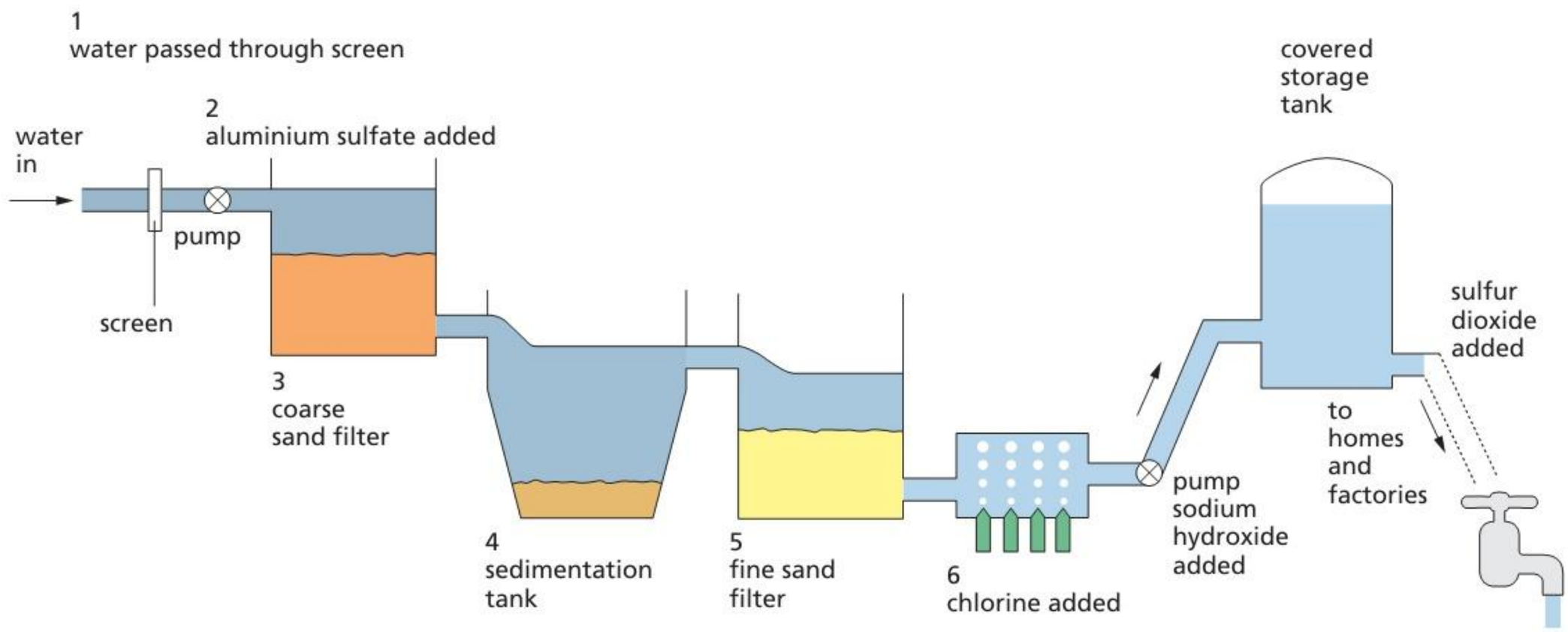


Figure 11.39 The processes involved in water treatment.

6 Finally, a little chlorine gas is added, which sterilises the water and kills any remaining bacteria. Excess chlorine can be removed by the addition of sulfur dioxide gas. The addition of chlorine gas makes the water more acidic and so appropriate amounts of sodium hydroxide solution are added. Fluoride is sometimes added to water if there is insufficient occurring naturally, as it helps to prevent tooth decay.

The 'iron problem'

If the acidity level of the treated water is not controlled, problems occur due to the precipitation of iron(III) hydroxide. These include:

- vegetables turning brown
- tea having an inky appearance and a bitter taste
- clothes showing rusty stains after washing (Figure 11.40).

Figure 11.40 The rusty stains on this pillowcase are due to iron (III) compounds in the water.

Sewage treatment

After we have used water, it must be treated again before it can be returned to rivers, lakes and seas. This multi-stage process known as sewage treatment is shown in Figure 11.41.

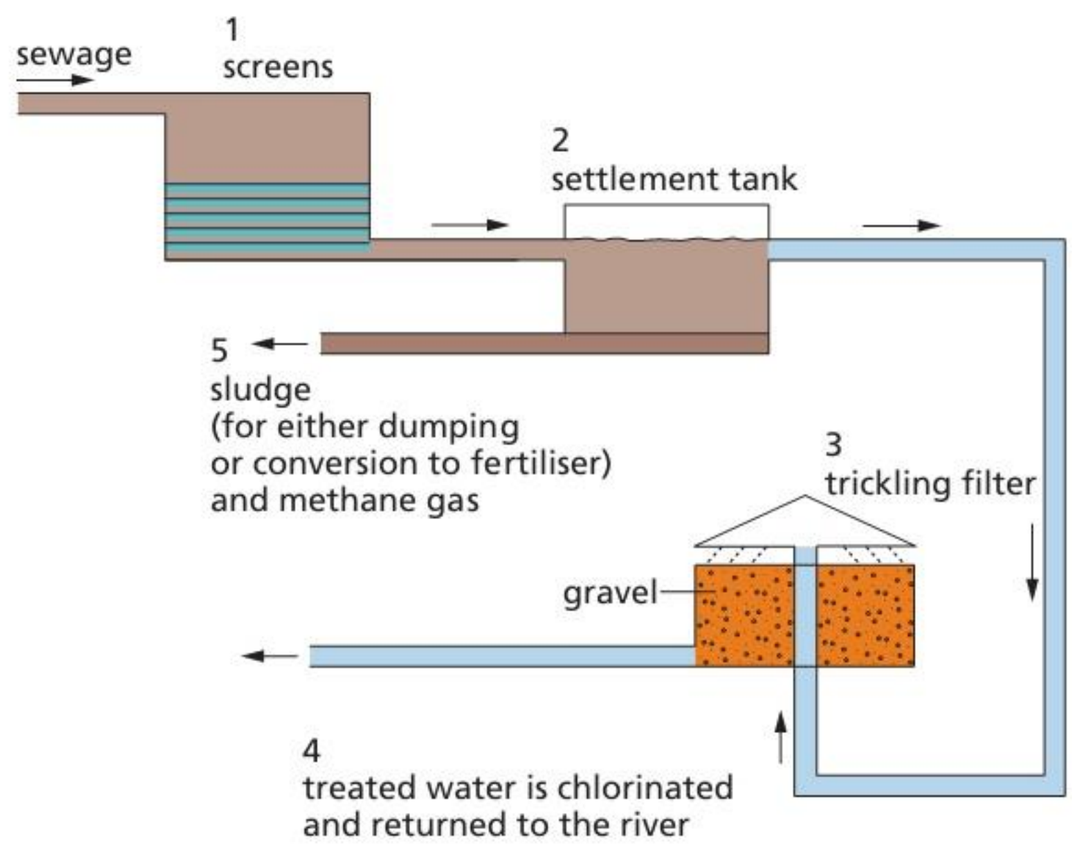


Figure 11.41 The processes involved in sewage treatment.

Used water, sewage, contains waste products such as human waste and washing-up debris as well as everything else that we put down a drain or sink. The processes that are involved in its treatment are as follows.

1 Large screens remove large pieces of rubbish.
2 Sand and grit are separated in large sedimentation tanks. The process is speeded up by adding aluminium sulfate, which helps the solids to

coagulate into larger particles that separate more rapidly. The sand and grit often contain large amounts of useful chemicals which, by the action of selected microbes, can be used as fertilisers.

3 The impure water is then removed and sent to a trickling filter, where it is allowed to drain through gravel on which microbes have been deposited. These kill off any remaining bacteria in the water by aerobic processes. This stage is known as biological filtration.

4 The treated water is then chlorinated and returned to a river, after checking.

5 Anaerobic bacteria digest what remains from the other stages. Methane gas is produced, which can be used as a fuel.

Water is a renewable resource but care must be taken to avoid excessive water pollution, as this could cause serious health issues such as typhoid or cholera.

Questions

1 Make a list of four major water pollutants and explain where they come from. What damage can these pollutants do?

2 In the treatment of water for public use, state the purpose of the addition of:
- a aluminium sulfate
- b chlorine
- c sodium hydroxide
- d sulfur dioxide.

3 Many industries use water as a coolant. Suggest the sorts of problems that may be created by this 'thermal pollution'.

A summary of the main sources of atmospheric and water pollution is shown in Table 11.9.

Table 11.9 Main sources of atmospheric and water pollution.

Pollutant	Source	Effects
Carbon dioxide, CO_2	Complete combustion of fossil fuels (p. 92) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$	Contributes to global warming (p. 212)
Carbon monoxide, CO	Incomplete combustion of fossil fuels (p. 95)	Poisonous, colourless, odourless gas
Oxides of nitrogen, NO_x	Car exhaust fumes (p. 182) $N_2(g) + O_2(g) \rightarrow 2NO(g)$ $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$	Oxides of nitrogen cause photochemical smog
Particulates, carbon C	Incomplete combustion of fossil fuels (p. 95)	Covers surfaces with carbon particles
Lead and lead oxides	Car batteries	Lead and its compounds are toxic and particularly affect the nervous system
Lead compounds	Lead additives are added to petrol (use is decreasing), resulting in exhaust fumes	Lead and its compounds are toxic
Chlorofluorocarbons (CFCs)	Aerosol sprays and refrigerators (p. 221)	Destroy the layer of ozone around the Earth
Nitrates NO_3^- and phosphates PO_4^{3-}	Fertilisers (p. 180)	Eutrophication*
Sodium phosphate	Detergents and water softeners	Eutrophication*
DDT*	Insecticides	Accumulates in food chain till eventually kills top predators
Paraquat* and glyphosate*	Herbicides	Kill harmless plants

*DDT is an organochlorine herbicide. It has high toxicity to insects but has low toxicity to mammals. However, it is very chemically stable so it accumulates in the environment and can become more concentrated in food chains. For this reason it is banned in many countries. However, as it is cheap many developing countries continue to use it.

*Paraquat is a total herbicide. It destroys all green plant material. Paraquat is inactivated as soon as it touches the soil but it does kill any plant whose leaves it touches.

*Glyphosate is a herbicide sprayed onto the outside of leaves and it then spreads through the plant to the roots. It is non-selective and kills useful plants as well as weeds.

*Eutrophication occurs when fertiliser drains into lakes and rivers, causing algae to multiply rapidly and the water to turn green. It results in fish and other organisms suffocating and dying through lack of oxygen in water.

Checklist

After studying Chapter 11 you should know and understand the following terms.

- **Artificial fertiliser** A substance added to soil to increase the amount of elements such as nitrogen, potassium and phosphorus. This enables crops grown in the soil to grow more healthily and to produce higher yields.
- **Atmosphere (air)** The mixture of gases that surrounds the Earth.
- **Bulk chemical** A chemical that, because of its large usage across a range of uses, is produced in large quantities.
- **Carbon dioxide** A colourless, odourless gas, soluble in water, producing a weak acid called carbonic acid. It makes up 0.04% of air. It is produced by respiration in all living things and by the burning of fossil fuels. It is taken in by plants in photosynthesis.
- **Chemical equilibrium** A dynamic state. The concentrations of the reactants and products remain constant because the rate at which the forward reaction occurs is the same as that of the back reaction.
- **Eutrophication** A process that occurs when fertiliser drains into lakes and rivers, causing algae to multiply rapidly and the water to turn green. it results in fish and other organisms suffocating and dying through lack of oxygen in the water.
- **Flue gas desulfurisation (FGD)** The process by which sulfur dioxide gas is removed from the waste gases of power stations by passing them through calcium hydroxide slurry.
- **Fractional distillation of air** The process used to extract individual gases from the air. Air is a major raw material. The mixture of gases is separated by first liquefying the mixture at low temperature and high pressure. The temperature is then allowed to rise and the gases collected as they boil off. The gases so produced have many and varied uses.
- **Haber process** The chemical process by which ammonia is made in very large quantities from nitrogen and hydrogen.
- **Hardness of water** This is caused by the presence of calcium (or magnesium) ions in water, which form a 'scum' with soap and prevent the formation of a lather. There are two types of hardness:
  - temporary hardness – caused by the presence of dissolved calcium (or magnesium) hydrogencarbonate
  - permanent hardness – this results mainly from dissolved calcium (or magnesium) sulfate.
- **Hydrogen bond** A weak intermolecular force (bond) that occurs between water molecules because the bonds within the molecules are polar. It is the strongest of these weak forces.
- **Limestone** A form of calcium carbonate ($CaCO_3$). Other forms include chalk, calcite and marble.
- **Nitrogen cycle** The system by which nitrogen and its compounds, both in the air and in the soil, are interchanged.
- **Nitrogen fixation** The direct use of atmospheric nitrogen in the formation of important compounds of nitrogen. Bacteria present in root nodules of certain plants are able to take nitrogen directly from the atmosphere to form essential protein molecules.
- **Optimum temperature** A compromise temperature used in industry to ensure that the yield of product and the rate at which it is produced make the process as economical as possible.
- **Ozone (trioxygen)** A colourless gas produced in the stratosphere by the action of high-energy ultraviolet radiation on oxygen gas, producing oxygen atoms. These oxygen atoms then react with further oxygen molecules to produce ozone. Its presence in the stratosphere acts as a screen (ozone layer) against dangerous ultraviolet radiation.
- **Particulates** Very small particles found in the atmosphere, such as certain types of smoke emitted from diesel engines, as well as dust.
- **Photosynthesis** The chemical process by which green plants synthesise their carbon compounds from atmospheric carbon dioxide using light as the energy source and chlorophyll as the catalyst.
- **Pollution** The modification of the environment caused by human influence. It often renders the environment harmful and unpleasant to life. Atmospheric pollution is caused by gases such as sulfur dioxide, carbon monoxide and nitrogen oxides being released into the atmosphere by a variety of industries and also by the burning of fossil fuels. Water pollution is caused by many substances, such as those found in fertilisers and in industrial effluent.
- **Raw material** A basic material from which a product is made. For example, the raw materials for the Haber process are nitrogen and hydrogen.
- **Removal of hardness** Temporary hardness is removed by boiling. Both temporary and permanent hardness are removed by:
  - addition of washing soda (sodium carbonate)
  - ion exchange
  - distillation.
- **Reversible reaction** A chemical reaction which can go both ways. This means that once some of the products have been formed they will undergo a chemical change once more to re-form the reactants. The reaction from left to right, as the equation for the reaction is written, is known as the forward reaction and the reaction from right to left is known as the back reaction.
- **Stratosphere** A layer of the atmosphere above the troposphere in which the ozone layer exists.
- **Transpiration** The process of evaporation of water from leaves.
- **Troposphere** A layer of the atmosphere closest to the Earth which contains about 75% of the mass of the atmosphere. The composition of dry air is relatively constant in this layer of the atmosphere.
- **Water cycle** The cycle by which water circulates around the Earth. The driving force behind the water cycle is the Sun.

Air and water

Additional questions

1 The apparatus shown on p. 174, Figure 11.6, was used to estimate the proportion of oxygen in the atmosphere.

A volume of dry air (200 cm^3) was passed backwards and forwards over heated copper until no further change in volume took place. The apparatus was then allowed to cool down to room temperature and the final volume reading was then taken. Some typical results are shown below.

Volume of gas before = 200 cm^3
Volume of gas after = 157 cm^3

During the experiment the copper slowly turned black.

a Why was the apparatus allowed to cool back to room temperature before the final volume reading was taken?
b Using the information given above, calculate the percentage volume reduction which has taken place.
c Explain briefly why there is a change in volume.
d What observation given above supports your explanation in **c**? Write a balanced chemical equation for any reaction which has occurred.
e Give the name of the main residual gas at the end of the experiment.
f Would you expect the copper to have increased or decreased in mass during the experiment? Explain your answer.

2 a Oxygen has an atomic number of 8 and mass number 16. What is the electron structure of the oxygen atom?
b How many electrons, neutrons and protons are there in the oxygen atom?
c Oxygen molecules are diatomic. Explain the meaning of this term.
d Draw a diagram of the oxygen molecule showing the outer shell of electrons only. What type of bonding does the molecule contain?

3 Explain the following.
a Air is a mixture of elements and compounds.
b The percentage of carbon dioxide in the atmosphere does not significantly vary from 0.04%.
c Power stations are thought to be a major cause of acid rain.

4 Use the words below to complete the following passage about limestone and its uses.

abundant **chlorine** **nitrates**
monitored **coolant** **fertilisers** **reactions**
pure **clay** **solvent** **essential**

Water is _____ for all life on Earth. It is the most _____ substance on Earth. Water is not only used for drinking and washing but also:

- in chemical _____ , for example in the production of ethanol from ethene
- as a _____ to dissolve things
- as a _____ in some chemical processes.

For us to drink it, water must be relatively _____. To make the water drinkable, any undissolved solids are removed through filtration beds. Aluminium sulfate is added to remove small particles of _____ . _____ is added to kill bacteria. Water is continuously _____ because there are certain substances that eventually find their way into the water supply. For example, artificial _____ are washed off farmland and this can cause the water supply to be contaminated with nitrates. _____ are harmful to babies.

5 a Explain what is meant by the term 'pollution' with reference to air and water.
b (i) Name an air pollutant produced by the burning of coal.
(ii) Name a different air pollutant produced by the combustion of petrol in a car engine.
c Some of our drinking water is obtained by purifying river water.
(i) Would distillation or filtration produce the purest water from river water? Give a reason for your answer.
(ii) Which process, distillation or filtration, is actually used to produce drinking water from river water? Comment on your answer in comparison to your answer in **c (i)**.
d Power stations produce warm water. This causes thermal pollution as this warm water is pumped into nearby rivers.
(i) Why do power stations produce such large quantities of warm water?
(ii) What effect does this warm water have on aquatic life?

6

Fertiliser	Formula	% nitrogen
Ammonia solution	NH_3	82.4
Calcium nitrate	$Ca(NO_3)_2$	17.1
Nitram	NH_4NO_3	35.0
Sodium nitrate	$NaNO_3$	
Potassium nitrate	KNO_3	

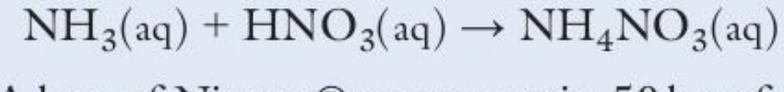
a Copy and complete the above table by calculating the percentage of nitrogen in the fertilisers sodium nitrate and potassium nitrate. (A_r: H = 1; N = 14; O = 16; Na = 23; K = 39; Ca = 40)

b Including the data you have just calculated, which of the fertilisers contains:
(i) the largest percentage of nitrogen?
(ii) the smallest percentage of nitrogen?

c Give the chemical name for the fertiliser that goes by the name Nitram®.

d Ammonia can be used directly as a fertiliser but not very commonly. Think of two reasons why ammonia is not often used directly as a fertiliser.

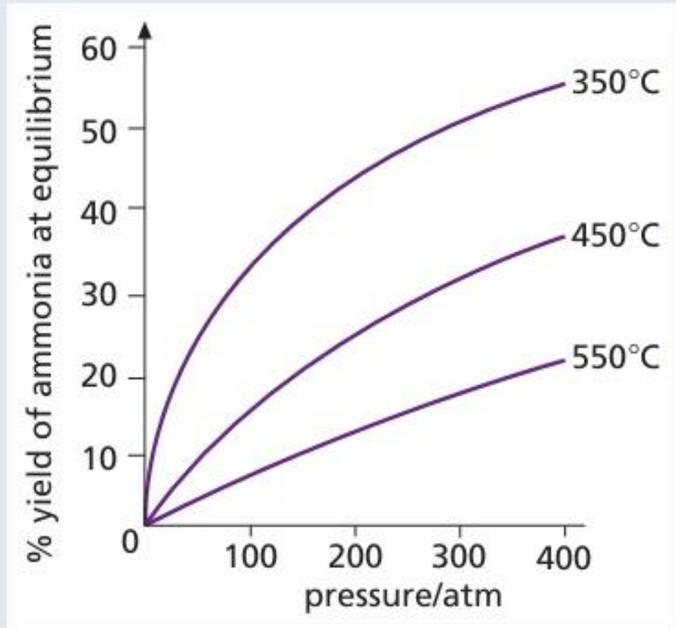
e Nitram® fertiliser is manufactured by the reaction of nitric acid with ammonia solution according to the equation:



$$NH_3(aq) + HNO_3(aq) \rightarrow NH_4NO_3(aq)$$

A bag of Nitram® may contain 50 kg of ammonium nitrate. What mass of nitric acid would be required to make it?

7 Ammonia gas is made industrially by the Haber process, which involves the reaction between the gases nitrogen and hydrogen. The amount of ammonia gas produced from this reaction is affected by both the temperature and the pressure at which the process is run. The graph shows how the amount of ammonia produced from the reaction changes with both temperature and pressure. The percentage yield of ammonia indicates the percentage of the nitrogen and hydrogen gases that are actually changed into ammonia gas.



a Write a word and balanced chemical equation for the reversible reaction between nitrogen and hydrogen to produce ammonia using the Haber process.

b What is meant by the term 'reversible reaction'?

c Use the graphs to say whether more ammonia is produced at:
(i) higher or lower temperatures
(ii) higher or lower pressures.

d What is the percentage yield of ammonia if the conditions used to run the process are:
(i) a temperature of 350 °C and a pressure of 100 atmospheres?
(ii) a temperature of 550 °C and a pressure of 350 atmospheres?

e The conditions in industry for the production of ammonia are commonly of the order of 200 atmospheres and 450 °C. What is the percentage yield of ammonia using these conditions?

f Why does industry use the conditions stated in part **e** if it is possible to obtain a higher yield of ammonia using different conditions?

8 Explain the following.

a Industry normally requires water which has been softened.

b Hard water is good for the promotion of healthy bones and teeth.

c Hard water causes kettles to fur. This 'fur' can be removed by using a dilute acid.

d Hard water wastes soap.

e Hard water can coat lead pipes and reduce the possibility of lead poisoning.

9 The results of testing five samples of water from different areas are shown in the table below. The soap solution was gradually added to 25 cm^3 of each sample of water with shaking until a permanent lather (one which lasts for 20 seconds) was obtained.

Water sample/ 25 cm^3	Volume of soap solution added/cm^3	
	Before boiling	After boiling
A	12	1
B	13	6
C	11	11
D	14	3
E	16	16

a **(i)** Which samples are permanently hard?
(ii) Which samples are temporarily hard?
(iii) Which sample contains both temporary and permanent hardness?

b Name a compound which could be present in sample D but not in sample E.

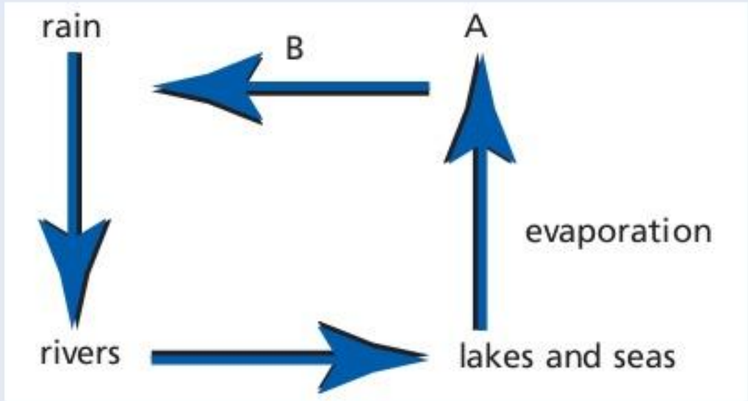
c Name a compound which could be present in sample E but not in sample D.

d Explain how the compound you have named in **c** gets into the water.

e Sample E was distilled. The water collected was tested with soap solution. What volume of soap solution might you expect to be added to produce a permanent lather? Comment on your answer.

10 A town's water supply is obtained by purifying river water.

a What are the **A** and **B** in the diagram of the water cycle below?



b **(i)** Filtration or distillation – which of these processes would produce the purest water from river water? Give a reason for your answer.
(ii) Filtration or distillation – which of these processes is used in this country for producing drinking water from river water?

c In recent years pollution of rivers and lakes has become a serious problem. State two main sources of river pollution.

d In some areas water is 'recycled'. What does recycling mean and why is it necessary?

12 Sulfur

Sulfur – the element
Uses of sulfur
Chemical properties of sulfur

Sulfur dioxide
Acid rain

Sulfuric acid
Industrial manufacture of sulfuric acid – the Contact process
Uses of sulfuric acid
Properties of sulfuric acid

Checklist

Additional questions

Sulfur – the element

Sulfur is a yellow non-metallic element. It is found in Group VI of the Periodic Table. It is a brittle, non-conducting solid with a fairly low melting point (115 °C). Sulfur will not dissolve in water but will dissolve in solvents such as carbon disulfide and methylbenzene (toluene). Like carbon, sulfur has allotropes. Its main allotropes are called rhombic sulfur and monoclinic sulfur (Figure 12.1).

Figure 12.1 Sulfur – rhombic (top) and monoclinic.

Sulfur is found in large quantities but in various forms throughout the world. It is found in metal ores such as copper pyrites or chalcopyrite ($CuFeS_2$) and zinc blende (ZnS) and in volcanic regions of the world. Natural gas and oil contain sulfur and its compounds, but the majority of this sulfur is removed as it would cause environmental problems. Sulfur obtained from these sources is known as 'recovered sulfur' and it is an important source of the element. It is also found as elemental sulfur in sulfur beds in Poland, Russia and the US (Louisiana). These sulfur beds are typically 200 m below the ground. Sulfur from these beds is extracted using the **Frasch process**, named after its inventor Hermann Frasch.

Uses of sulfur

Sulfur has a very important role in the chemical industry. The vast majority of sulfur is used to produce perhaps the most important industrial chemical, sulfuric acid. Sulfur is also used to **vulcanise** rubber, a process which makes the rubber harder and increases its elasticity. Relatively small amounts are used in the manufacture of matches, fireworks and fungicides, as a sterilising agent and in medicines.

Chemical properties of sulfur

Sulfur will react with both metals and non-metals.

- It reacts with magnesium metal to form magnesium sulfide.

$$\text{magnesium} + \text{sulfur} \rightarrow \text{magnesium sulfide}$$
$$Mg(s) + S(s) \rightarrow MgS(s)$$

- It reacts with oxygen to produce sulfur dioxide gas.

$$\text{sulfur} + \text{oxygen} \rightarrow \text{sulfur dioxide}$$
$$S(s) + O_2(g) \rightarrow SO_2(g)$$

Sulfur dioxide

Sulfur dioxide is a colourless gas produced when sulfur or substances containing sulfur, for example crude oil or natural gas, are burned in oxygen gas. It has a choking smell and is extremely poisonous. The gas dissolves in water to produce an acidic solution of sulfurous acid.

sulfur dioxide + water ⇌ sulfurous acid

$$SO_2(g) + H_2O(l) \rightleftharpoons H_2SO_3(aq)$$

This solution turns moist universal indicator paper red. It is one of the major pollutant gases and is the gas principally responsible for **acid rain**. However, it does have some uses: as a bleaching agent for paper manufacture, in fumigants and in the preservation of food by killing bacteria.

Sulfur dioxide will turn acidified potassium manganate(VII) solution from purple to colourless (Figure 12.2).

Figure 12.2 Test for sulfur dioxide.

Question

1 'Sulfur is a non-metallic element.' Discuss this statement, giving physical and chemical reasons to support your answer.

Acid rain

Rainwater is naturally acidic since it dissolves carbon dioxide gas from the atmosphere as it falls. Natural rainwater has a pH of about 5.7. In recent years, especially in central Europe, the pH of rainwater has fallen to between pH 3 and pH 4.8. This increase

Figure 12.3 This forest has been devastated by acid rain.

in acidity has led to extensive damage to forests (Figure 12.3), lakes and marine life.

In addition it has led to the increased corrosion of exposed metals and to damage to buildings and statues made from limestone or marble (Figure 12.4). The sulfurous acid in rainwater oxidises to sulfuric acid. The sulfuric acid reacts with the limestone, which is eaten away by the chemical process.

limestone + sulfuric acid → calcium sulfate + water + carbon dioxide

$$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + H_2O(l) + CO_2(g)$$

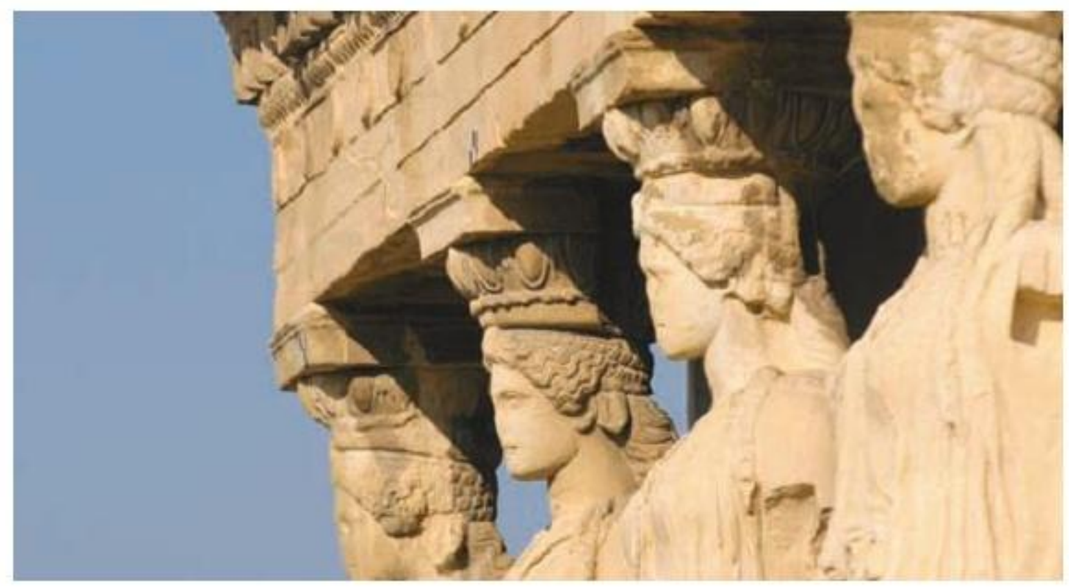


Figure 12.4 Acid rain is responsible for much of the damage to this temple on the Acropolis in Athens.

The amount of sulfur dioxide in the atmosphere has increased dramatically over recent years. There has always been some sulfur dioxide in the atmosphere, from natural processes such as volcanoes and rotting vegetation. Over Europe, however, around 80% of the sulfur dioxide in the atmosphere is formed from the combustion of fuels containing sulfur (Figure 12.5). After dissolving in rain to produce sulfurous acid, it reacts further with oxygen to form sulfuric acid.

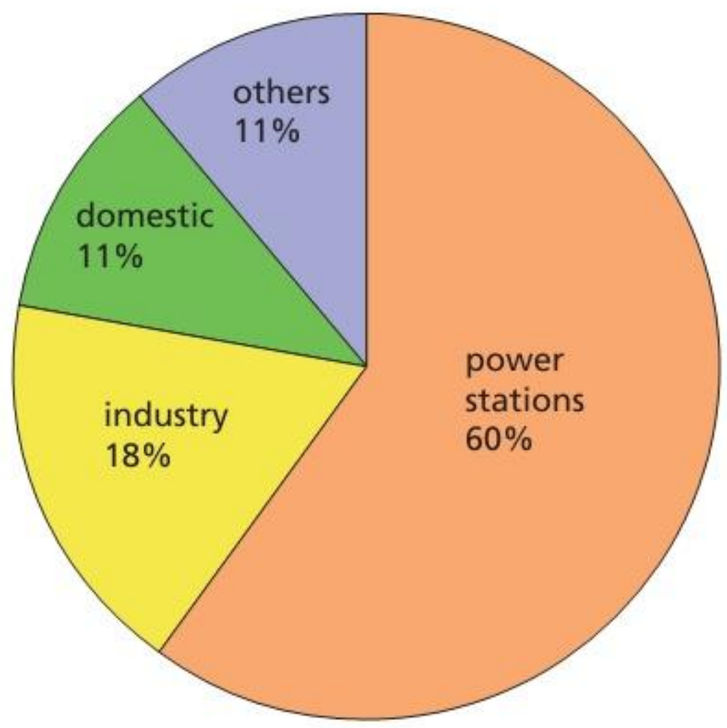


Figure 12.5 Sources of sulfur dioxide.

Nitric acid (HNO_3) in the atmosphere, formed from NO_2 gas dissolving in the rainwater, also contributes to the production of acid rain.

Questions

1 How could the amount of sulfur dioxide being produced by the above sources be reduced?

2 Devise an experiment which you could carry out in the school laboratory to determine the amount of sulfur in two different types of coal.

Sulfuric acid

Industrial manufacture of sulfuric acid – the Contact process

The major use of sulfur is in the production of sulfuric acid. This is probably the most important industrial chemical, and the quantity of it produced by a country has been linked with the economic stability of the country. In excess of 150 million tonnes of sulfuric acid are produced worldwide each year. It is used mainly as the raw material for the production of many substances (Figure 12.6).

The process by which sulfuric acid is produced is known as the **Contact process** (Figure 12.7).

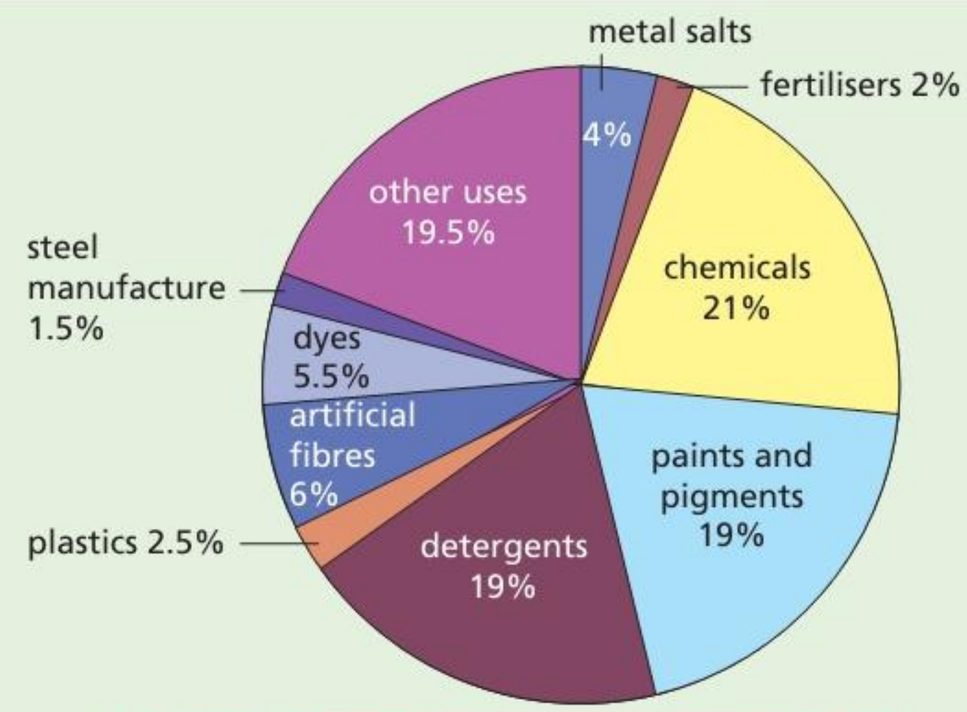


Figure 12.6 Products made from sulfuric acid include detergents, paints and pigments.

Figure 12.7 A Contact process plant used for making sulfuric acid.

The process has the following stages.

- Sulfur dioxide is first produced, primarily by the reaction of sulfur with air.

$$\text{sulfur} + \text{oxygen} \rightarrow \text{sulfur dioxide}$$
$$S(s) + O_2(g) \rightarrow SO_2(g)$$

- Any dust and impurities are removed from the sulfur dioxide produced, as well as any unreacted oxygen. These 'clean' gases are heated to a temperature of approximately 450 °C and fed into a reaction vessel, where they are passed over a catalyst of vanadium(V) oxide (V_2O_5). This catalyses the reaction between sulfur dioxide and oxygen to produce sulfur trioxide (sulfur(VI) oxide, SO_3).

$$\text{sulfur dioxide} + \text{oxygen} \rightleftharpoons \text{sulfur trioxide}$$
$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \quad \Delta H = -197\ \text{kJ mol}^{-1}$$

This reaction is reversible and so the ideas of Le Chatelier (Chapter 11, p. 178) can be used to increase the proportion of sulfur trioxide in the equilibrium mixture. The forward reaction is exothermic and so would be favoured by low temperatures. The temperature of 450 °C used is an optimum temperature which produces sufficient sulfur trioxide at an economical rate. Since the reaction from left to right is also accompanied by a decrease in the number of molecules of gas, it will be favoured by a high pressure. In reality, the process is run at atmospheric pressure. Under these conditions, about 96% of the sulfur dioxide and oxygen are converted into sulfur trioxide. The heat produced by this reaction is used to heat the incoming gases, thereby saving money.

- If this sulfur trioxide is added directly to water, sulfuric acid is produced. This reaction, however, is very violent and a thick mist is produced.

$$\text{sulfur trioxide} + \text{water} \rightarrow \text{sulfuric acid}$$
$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(l)$$

This acid mist is very difficult to deal with and so a different route to sulfuric acid is employed. Instead, the sulfur trioxide is dissolved in concentrated sulfuric acid (98%) to give a substance called **oleum**.

$$\text{sulfuric acid} + \text{sulfur trioxide} \rightarrow \text{oleum}$$
$$H_2SO_4(aq) + SO_3(g) \rightarrow H_2S_2O_7(l)$$

The oleum formed is then added to the correct amount of water to produce sulfuric acid of the required concentration.

$$\text{oleum} + \text{water} \rightarrow \text{sulfuric acid}$$
$$H_2S_2O_7(l) + H_2O(l) \rightarrow 2H_2SO_4(l)$$

Questions

1 Produce a flow diagram to show the different processes which occur during the production of sulfuric acid by the Contact process. Write balanced chemical equations showing the processes which occur at the different stages, giving the essential raw materials and conditions used.

2 Both the following reactions are reversible:
(i) $X_2(g) + O_2(g) \rightarrow 2XO(g)$
(ii) $2XO(g) + O_2(g) \rightarrow 2XO_2(g)$
Suggest a reason why an increase in pressure:
a does not favour reaction (i)
b increases the amount of XO_2 produced in reaction (ii).

Uses of sulfuric acid

Sulfuric acid has many uses in industry. It is such an important **bulk chemical** that the amount of sulfuric acid which a country uses in one year can be seen as a measure of that country's economic development, that is, how modern or wealthy it is.

For example, concentrated sulfuric acid is used:

- for making detergents (Chapter 15, p. 239)
- as a catalyst (Chapter 7, p. 109 and Chapter 14, p. 224)
- as a dehydrating agent (see p. 202 for the effect it has on hydrates such as copper(II) sulfate pentahydrate).

However, the concentrated acid is often mixed with water to form dilute sulfuric acid. The diluted acid is used:

- for making fertilisers (Chapter 11, p. 180)
- for treating metals to remove oxidation ('pickling') before painting
- in anodising aluminium (Chapter 5, p. 76)
- as the acid in car batteries
- for making paints, dyes and fibres
- as a common laboratory reagent.

Properties of sulfuric acid

Dilute sulfuric acid

Dilute sulfuric acid is a typical strong **dibasic** acid. A dibasic acid is one with two replaceable hydrogen atoms which may produce two series of salts – normal and acid salts (Chapter 8, pp. 122 and 126).

It will react with bases such as sodium hydroxide and copper(II) oxide to produce normal salts, called **sulfates**, and water.

- With sodium hydroxide:

sodium hydroxide	+	sulfuric acid	→	sodium sulfate	+	water

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$$

- With copper(II) oxide:

copper(II) oxide	+	sulfuric acid	→	copper(II) sulfate	+	water

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

It also reacts with carbonates to give normal salts, carbon dioxide and water, and with reactive metals to give a normal salt and hydrogen gas. The reaction between zinc and sulfuric acid is often used to prepare hydrogen gas in the laboratory (Figure 12.8).

zinc	+	sulfuric acid	→	zinc sulfate	+	hydrogen

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$

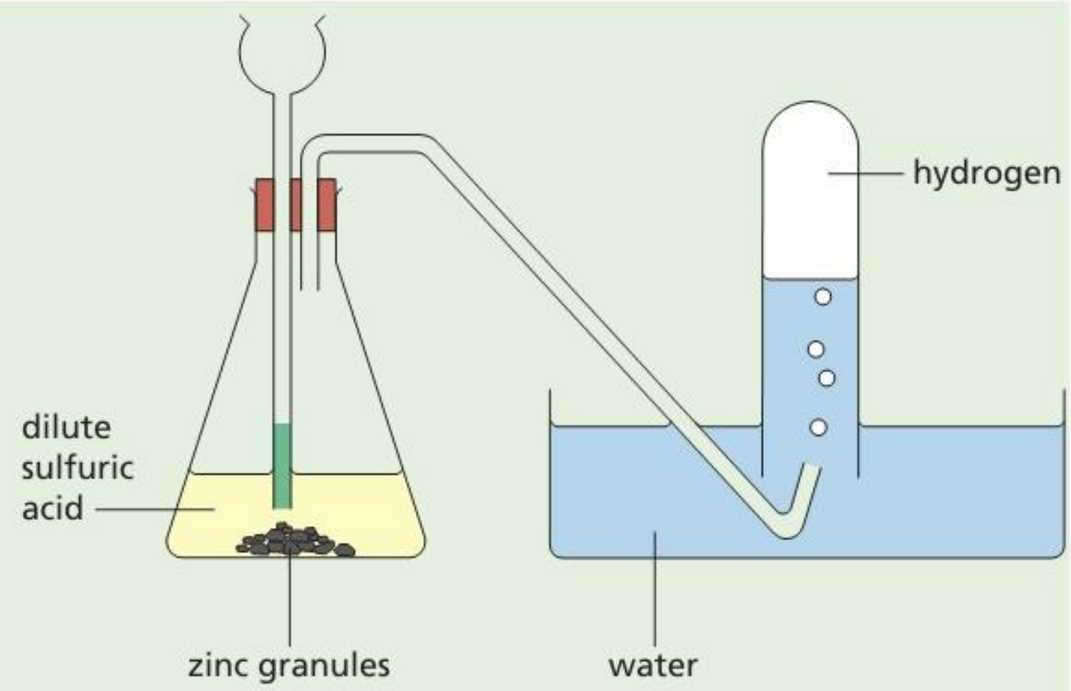


Figure 12.8 The laboratory preparation of hydrogen gas.

The preparation of the acid salt with sodium hydroxide requires twice the volume of acid as that used in the preparation of the normal salt. Therefore, if 25 cm^3 of dilute sulfuric acid were required to form the normal salt from a given volume of alkali of a particular concentration then 50 cm^3 of the same acid solution would be required to produce the acid salt, sodium hydrogensulfate, from the same volume of alkali.

sodium hydroxide	+	sulfuric acid	→	sodium hydrogensulfate	+	water

$$NaOH(aq) + H_2SO_4(aq) \rightarrow NaHSO_4(aq) + H_2O(l)$$

Sulfates

The salts of sulfuric acid, sulfates, can be identified by a simple test-tube reaction. To test for a sulfate, add a few drops of dilute hydrochloric acid to your unknown followed by a few drops of barium chloride. If a sulfate is present, a white precipitate of barium sulfate forms.

barium ions	+	sulfate ions	→	barium sulfate

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

Many sulfates have very important uses, as can be seen from Table 12.1.

Table 12.1 Uses of some metal sulfates.

Salt	Formula	Use
Ammonium sulfate	$(NH_4)_2SO_4$	Fertiliser
Barium sulfate	$BaSO_4$	'Barium meal' used in diagnostic medical X-ray studies
Calcium sulfate	$CaSO_4 \cdot \frac{1}{2}H_2O$	'Plaster of Paris' used to set bones
Magnesium sulfate	$MgSO_4$	In medicine it is used as a laxative

Concentrated sulfuric acid

Concentrated sulfuric acid is a powerful **dehydrating agent** – it will take water from a variety of substances. One such substance is cane sugar, or sucrose (Figure 12.9).

sucrose (sugar) $\xrightarrow{\text{conc } H_2SO_4}$ carbon + water

$C_{12}H_{22}O_{11}(s) \longrightarrow 12C(s) + 11H_2O(l)$

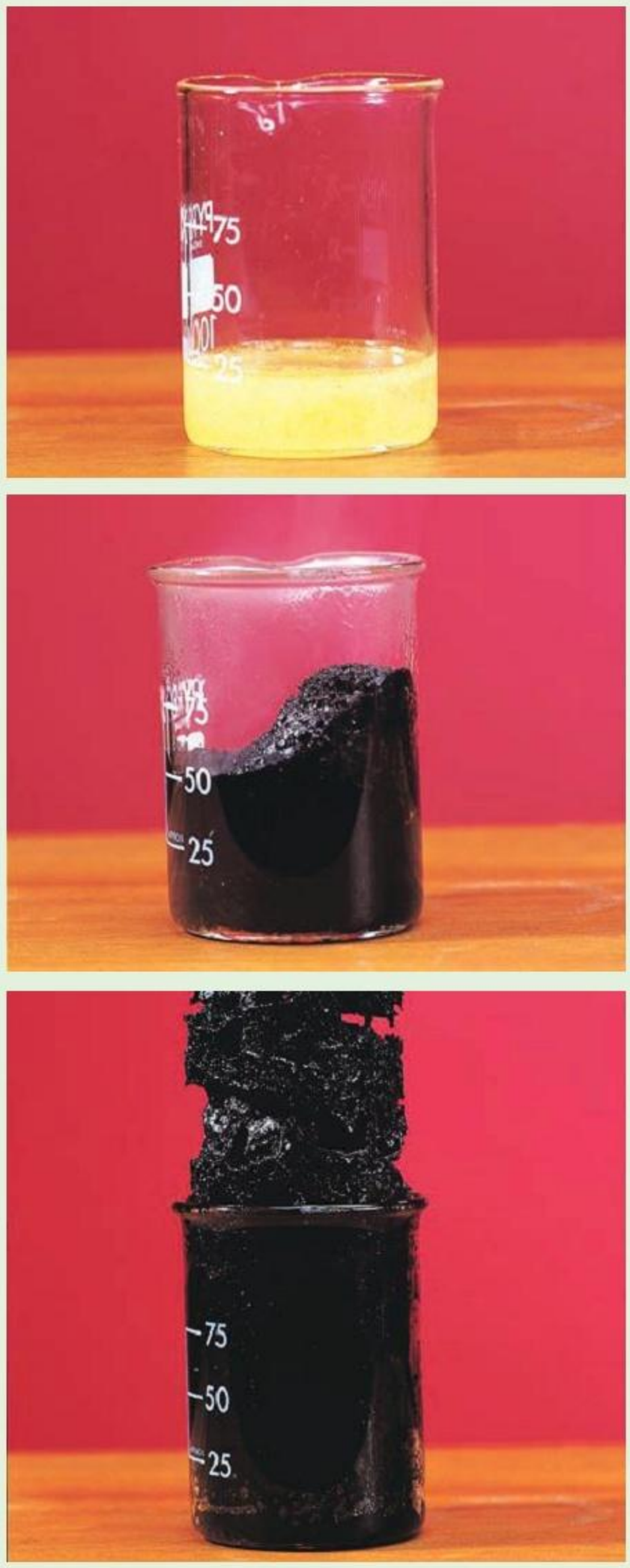


Figure 12.9 The concentrated sulfuric acid has removed the elements of water from the sugar, leaving black carbon.

Concentrated sulfuric acid will also take water from hydrated copper(II) sulfate crystals, leaving only anhydrous copper(II) sulfate. If a few drops of concentrated sulfuric acid are added to some blue hydrated copper(II) sulfate crystals, they slowly turn white as the water of crystallisation is removed by the acid. Eventually, only a white powder – anhydrous copper(II) sulfate – remains.

hydrated copper(II) sulfate $\xrightarrow{\text{conc } H_2SO_4}$ anhydrous copper(II) sulfate + water

$CuSO_4.5H_2O(s) \longrightarrow CuSO_4(s) + 5H_2O(l)$

Concentrated sulfuric acid should be treated very carefully, because it will also remove water from flesh! It is a very corrosive substance and should always be handled with care.

Diluting concentrated sulfuric acid must be done with great care because of its affinity for water. The concentrated sulfuric acid should always be added to the water, not the other way around.

Many gases are also dried by passing them through concentrated sulfuric acid. One exception is ammonia. Ammonia is alkaline so it cannot be dried in this way as it would react with the sulfuric acid.

Concentrated sulfuric acid will also act as a powerful **oxidising agent**. For example, when heated with metals such as copper it will oxidise them to copper ions and form copper(II) sulfate.

copper + concentrated sulfuric acid $\xrightarrow{\text{heat}}$ copper(II) sulfate + water + sulfur dioxide

$Cu(s) + 2H_2SO_4(l) \longrightarrow CuSO_4(aq) + 2H_2O(l) + SO_2(g)$

This reaction is often used as a method of making sulfur dioxide in the laboratory.

An interesting property of concentrated sulfuric acid is that it can be used to make nitric acid and hydrochloric acid (both these acids are more volatile than concentrated sulfuric acid).

If concentrated sulfuric acid is added to a metal chloride such as sodium chloride, hydrogen chloride gas is produced.

sodium chloride + concentrated sulfuric acid → sodium hydrogensulfate + hydrogen chloride

$NaCl(s) + H_2SO_4(l) \rightarrow NaHSO_4(s) + HCl(g)$

If the hydrogen chloride is dissolved in water, then dilute hydrochloric acid is produced.

If concentrated sulfuric acid is heated with a metal nitrate such as sodium nitrate, nitric acid vapour is produced.

sodium nitrate + concentrated sulfuric acid → sodium hydrogensulfate + nitric acid

$$NaNO_3(s) + H_2SO_4(l) \rightarrow NaHSO_4(s) + HNO_3(g)$$

Neither of these methods would be suitable for making these acids on a large scale.

Questions

1 If you were given an unlabelled bottle which was thought to be dilute sulfuric acid, how would you show that the solution contained sulfate ions ($SO_4^{2-}(aq)$), how would you show that it was an acid and how would you determine the concentration of the acid?

2 Write balanced chemical equations for the reactions between dilute sulfuric acid and:
   a zinc oxide
   b potassium carbonate
   c aluminium.

3 Describe, with the aid of equations, how concentrated sulfuric acid can react as:
   a a dehydrating agent
   b an oxidising agent.

Checklist

After studying Chapter 12 you should know and understand the following terms.

- **Acid rain** Rainwater with a pH in the range 3 to 4.8.
- **Bulk chemical** A chemical produced in large quantities by standard chemical reactions, for example ammoina and sulfuric acid.
- **Contact process** The industrial manufacture of sulfuric acid using the raw materials sulfur and air.
- **Dehydrating agent** A substance that will remove water or the elements of water from another substance.
- **Dibasic acid** An acid which contains two replaceable hydrogen atoms per molecule of the acid, for example sulfuric acid, H_2SO_4.
- **Oleum** Very concentrated sulfuric acid (98%).
- **Soapless detergents** Substances which are more effective than soap at producing lathers, especially in hard water areas. They are large organic molecules, produced using sulfuric acid.
- **Sulfate** A salt of sulfuric acid formed by the reaction of the acid with carbonates, bases and some metals. It is possible to test for the presence of a sulfate by the addition of dilute hydrochloric acid and some barium chloride solution. A white precipitate of barium sulfate is formed if a sulfate is present.

Sulfur

Additional questions

1 Sulfur has an atomic number of 16 and it is represented as:

$^{32}_{16}S$

a How many electrons, protons and neutrons does a sulfur atom contain?
b How many electrons will there be in the outer shell (energy level) of this atom?
c To which group of the Periodic Table does sulfur belong?
d Write down the names of two other elements found in the same group as sulfur.
e Write down the formula of the ion that sulfur will form in reactions with metals such as magnesium.
f Sulfur will react with powdered calcium when the mixture is heated. Write a word and a balanced chemical equation for this reaction.
g Give two large-scale uses of sulfur.

2 Explain the following.
a Chemical plants that produce sulfuric acid are often located on the coast.
b Even though more sulfuric acid could be produced using high pressures, normal atmospheric pressure is used.
c Natural rubber cannot be used to produce car tyres but vulcanised rubber can.
d Sulfur dioxide gas is regarded as a pollutant.
e Coal-fired and oil-fired power stations produce sulfur dioxide. Some of them are being fitted with flue gas desulfurisation (FGD) units.

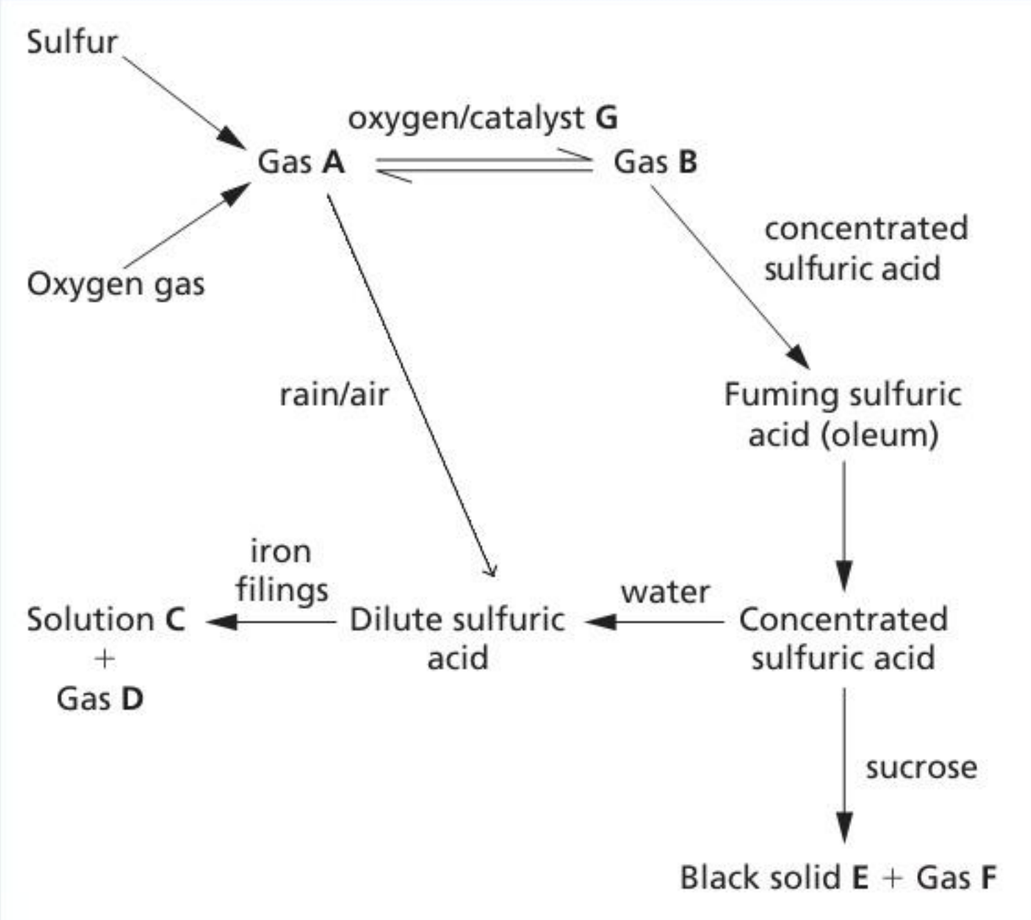
3 A type of coal contains 0.5% of sulfur by mass.
a Write an equation for the formation of sulfur dioxide gas when this coal is burned.
b What mass of sulfur is contained in 1500 tonnes of coal?
c What mass of sulfur dioxide gas would be formed if 1500 tonnes of coal were burned?
d What volume would this mass of sulfur dioxide gas occupy, measured at room temperature and pressure (rtp)? (A_r: O = 16; S = 32. One mole of a gas occupies 24 dm^3 at rtp.)

4 Fossil fuels, such as oil, coal and natural gas, all contain some sulfur. When these fuels are burned they produce many different gases. Concern has grown in recent years about the effects of one of these gases, sulfur dioxide. When sulfur dioxide dissolves in rainwater it forms an acidic solution which has become known as acid rain.
Money has been made available to solve the problem of acid rain. Attempts are being made to clean gases being released from power stations and to look into ways in which the effects of acid rain can be reversed. The table below and Figure 12.5 (p. 199) give some data about the emission of sulfur dioxide.

	Million tonnes per year
USA	26
Russia/Ukraine	18
Germany	7
UK	5
Canada	5
France	3
Poland	3
Italy	3
Other countries	30

a Using the figures in the table, produce a bar chart to show the amount of sulfur dioxide produced by each of the countries listed.
b What percentage of the world's sulfur dioxide is produced in:
(i) France?
(ii) North America?
c Using the information above, explain why countries such as the US, Russia and Germany are at the top of the list of sulfur dioxide producers.
d If the total amount of sulfur dioxide produced by Canada is 5 million tonnes per year, what amount is produced by:
(i) power stations?
(ii) domestic users?
(iii) industry?

5 Study the following reaction scheme:



a Identify the substances **A** to **F** by giving their names and formulae.
b Write a balanced chemical equation for the formation of gas **B**.
c **(i)** Describe a chemical test, and give the positive result of it, to identify gas **D**.
(ii) Describe a chemical test, and give the positive result of it, to identify gas **F**.
d How would you obtain solid **C** from the solution **C**?
e Which pathway shows the formation of acid rain?
f In which way is the concentrated sulfuric acid acting in its reaction with sucrose?
g Where does the oxygen gas come from to form gas **A**?

6 In a neutralisation experiment, 25 cm^3 of dilute sulfuric acid was required to react completely with 40 cm^3 of a solution of 0.25 mol dm^{-3} potassium hydroxide.
a Write a balanced chemical equation for the reaction between dilute sulfuric acid and potassium hydroxide.
b Calculate the number of moles of potassium hydroxide solution used in the reaction.
c How many moles of dilute sulfuric acid would this number of moles of potassium hydroxide react with?
d Calculate the concentration of the dilute sulfuric acid.
e Which indicator could have been used to determine when neutralisation had just occurred?

7 Describe how you would prepare some crystals of hydrated copper(II) sulfate from copper(II) oxide and dilute sulfuric acid. Draw a diagram of the apparatus you would use and write a balanced chemical equation for the reaction.

8 Sulfuric acid is manufactured by the Contact process, in which a mixture of sulfur dioxide and excess air, preheated to 450 °C, is passed over a catalyst.
a How is the sulfur dioxide made? Give an equation for the reaction.
b Write the balanced chemical equation for the reaction that takes place between sulfur dioxide and air.
c Name a suitable catalyst for the reaction.
d Why is a catalyst necessary for this process?
e The product of this reaction will react directly with water. When it does so, it forms sulfuric acid. However, the method is impractical because the reaction is exothermic and an acid mist forms. Describe the modern indirect method of changing the product of the reaction between sulfur dioxide and air to form sulfuric acid.
f Give two uses of concentrated sulfuric acid.

13 Inorganic carbon chemistry

Limestone
Direct uses of limestone
Indirect uses of limestone

Carbonates
Properties of carbonates

Carbon dioxide
Uses of carbon dioxide
Laboratory preparation of carbon dioxide gas
Properties of carbon dioxide gas

Checklist

Additional questions

Limestone

Environmental Problems in the Peak District National Park

Local residents have been protesting against proposals to site a new limestone quarry in the beautiful Derbyshire Peak District in the UK. However, limestone is such a useful and sought-after mineral that demand has encouraged mining in National Park areas.

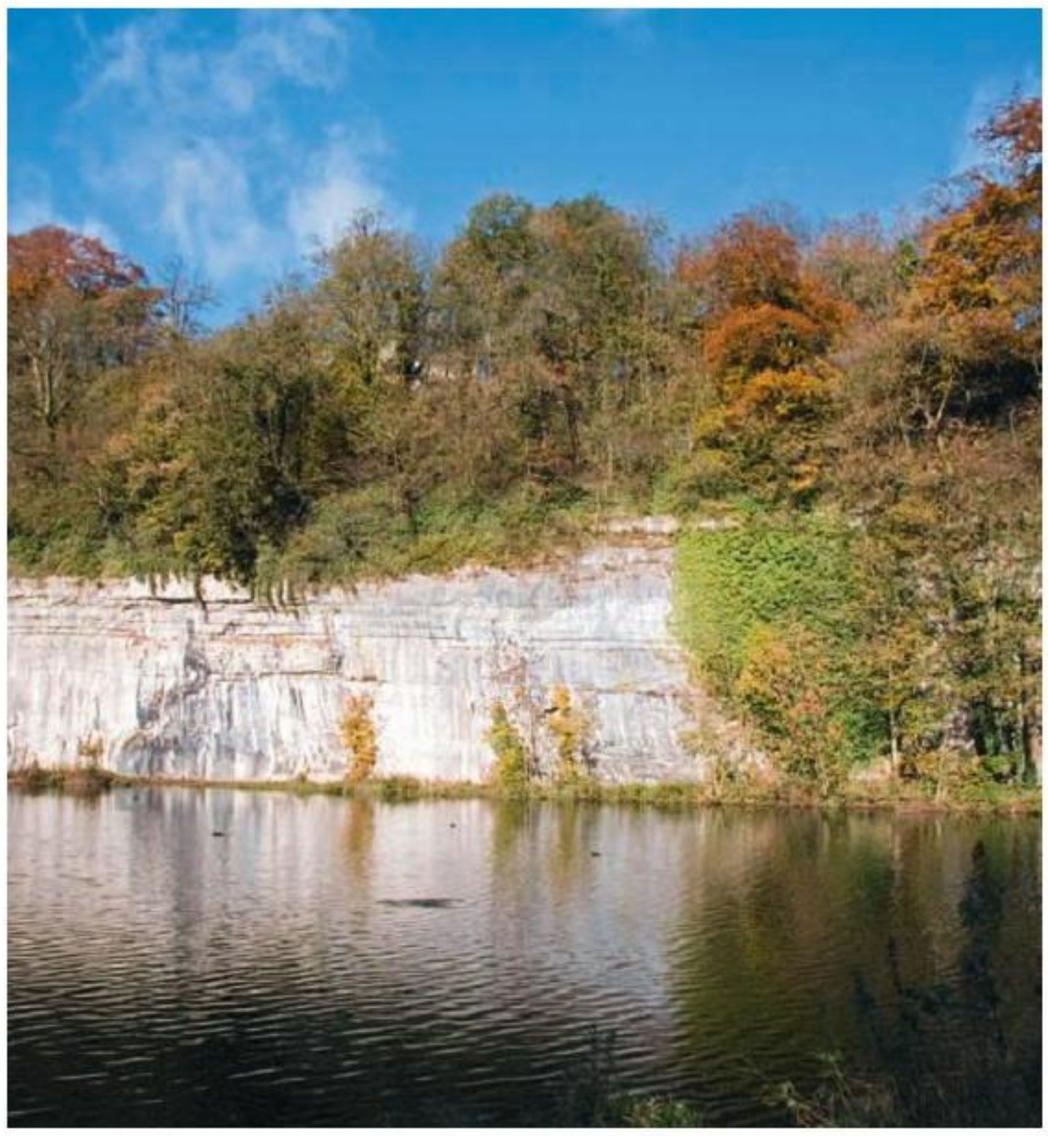


Figure 13.1 This gorge in the Peak District National Park in the UK is made from limestone.

As the newspaper article on the left says, limestone is found in the Peak District, an example being this gorge on the River Wye in the UK (Figure 13.1). As well as giving rise to beautiful and varied countryside, limestone is a very useful raw material.

Limestone is composed of calcium carbonate ($CaCO_3$) in the form of the mineral calcite (Figure 13.2). Chalk and marble are also made of calcite which is the second most abundant mineral in the Earth's crust after the different types of silicates (which include clay, granite and sandstone).

Chalk is made of the 'shells' of marine algae (that is, plants). It is a form of limestone. Most other limestones are formed from the debris of animal structures, for example brachiopods and crinoids.

Marble is a metamorphic rock made of calcium carbonate. It is formed when limestone is subjected to high pressures or high temperatures, or sometimes both acting together, to create crystals of calcium carbonate in the rock.

In a typical year, in excess of 500 million tonnes of limestone are quarried worldwide. Although it is cheap to quarry, as it is found near the surface, there are some environmental costs in its extraction.

Question

1 List the environmental issues which could arise through the quarrying of limestone.

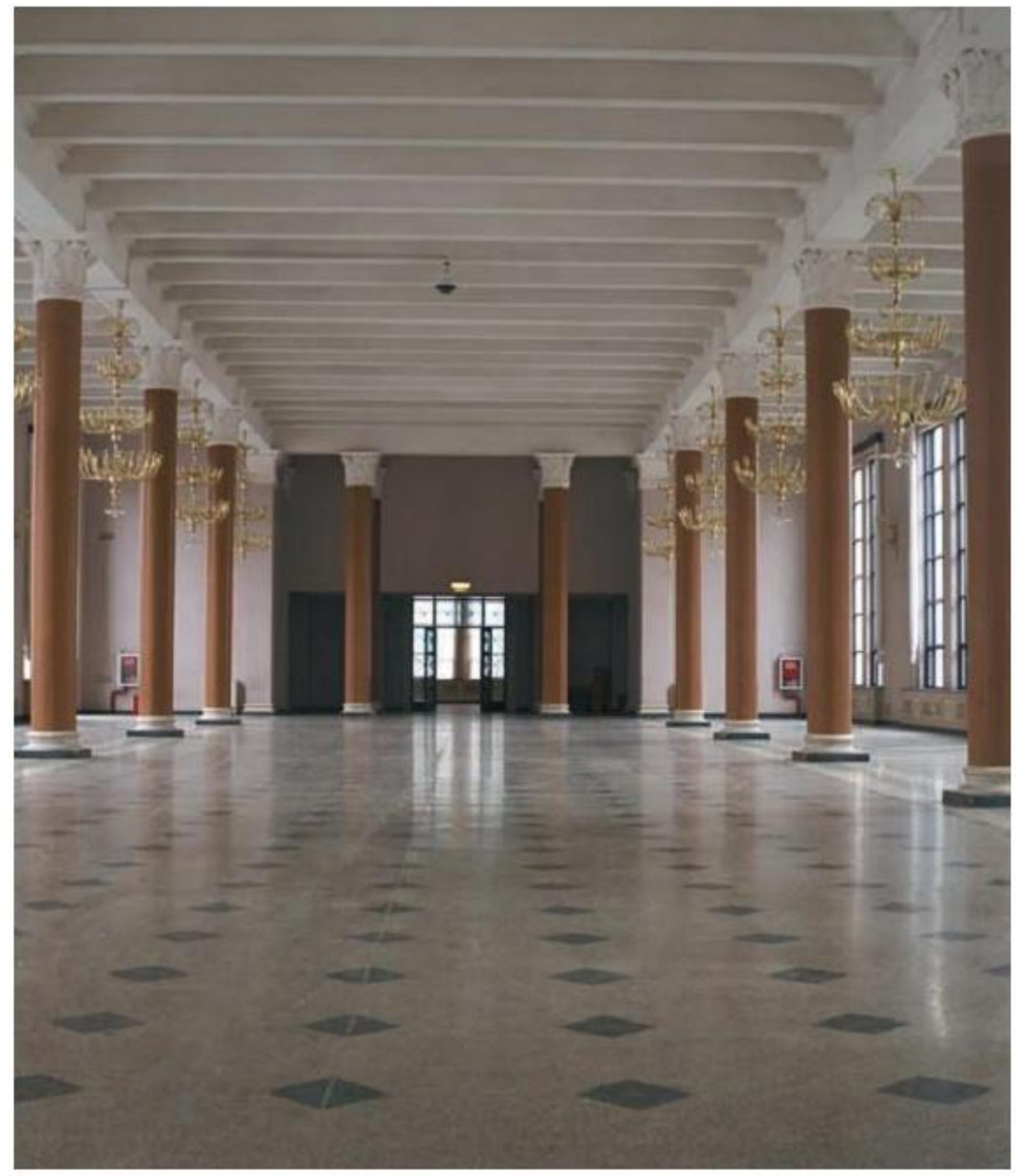


Figure 13.2 Chalk, calcite and marble are all forms of calcium carbonate.

Direct uses of limestone

Limestone has a variety of uses in, for example, the making of cement, road building, glass making and the extraction of iron (Figure 13.3).

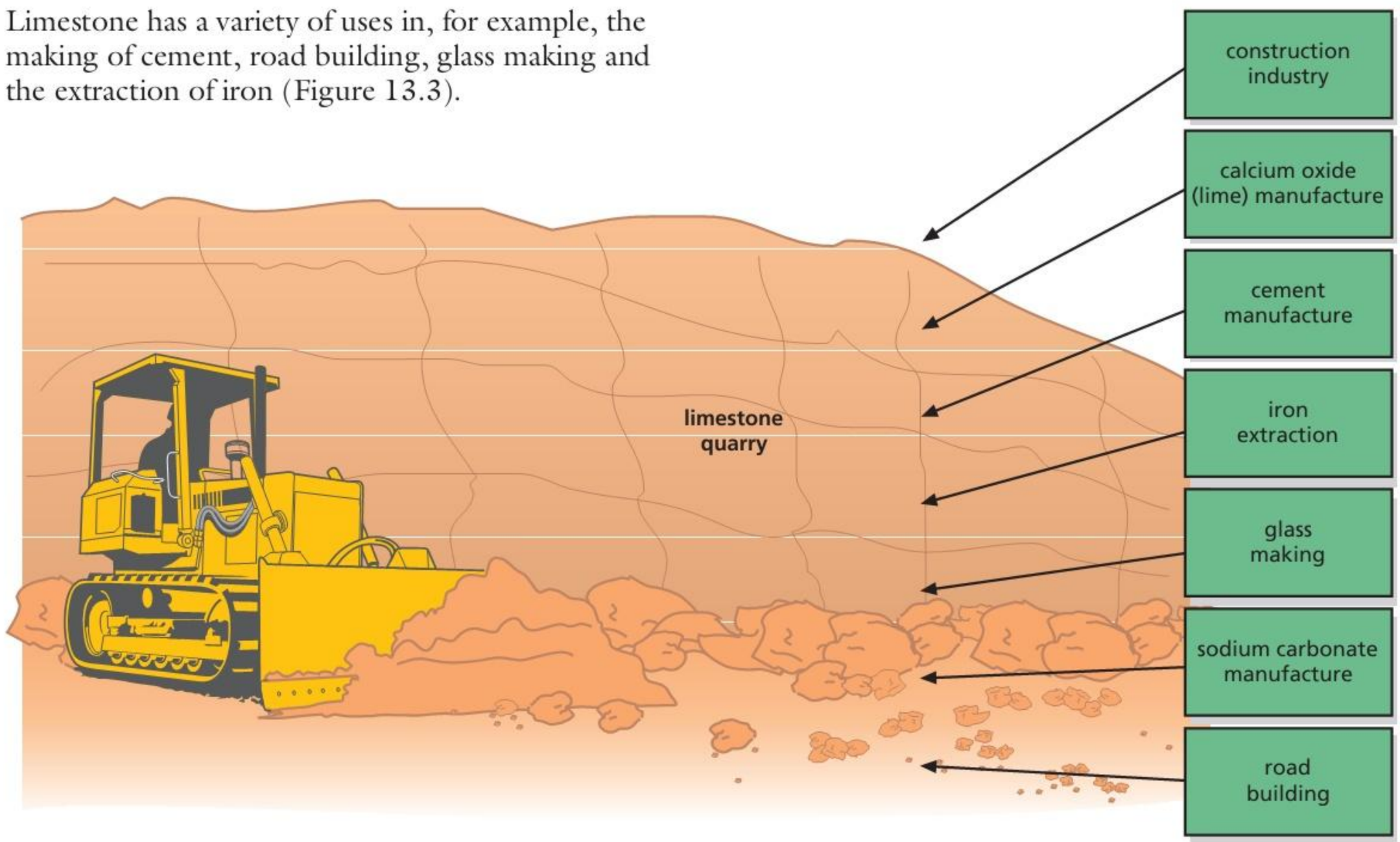


Figure 13.3 Uses of limestone.

Neutralisation of acid soil

Powdered limestone is most often used to neutralise acid soil (Figure 13.4) because it is cheaper than any form of lime (calcium oxide), which has to be produced by heating limestone (see p. 209), and because it is slow acting and an excess does not make the soil alkaline. The reaction of limestone with acidic soil can be shown by the following ionic equation.

carbonate ion + hydrogen ion → carbon dioxide + water

(from $CaCO_3$) (from acid soils)

$$CO_3^{2-}(s) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$$

Figure 13.4 Spreading limestone on to soil.

Manufacture of iron and steel

In the blast furnace, limestone is used to remove earthy and sandy materials found in the iron ore. A liquid slag is formed, which is mainly calcium silicate. More details of the extraction of iron and its conversion into steel are given in Chapter 10.

Manufacture of cement and concrete

Limestone (or chalk) is mixed with clay (or shale) in a heated rotary kiln, using coal or oil as the fuel (Figure 13.5). The material produced is called

Figure 13.5 A modern rotary kiln.

cement. It contains a mixture of calcium aluminate ($Ca(AlO_2)_2$) and calcium silicate ($CaSiO_3$).
The dry product is ground to a powder and then a little calcium sulfate ($CaSO_4$) is added to slow down the setting rate of the cement. When water is added to the mixture, slow complex chemical changes occur, resulting in the formation of a hard interlocking mass of crystals of hydrated calcium aluminate and silicate.

Concrete is a mixture of cement with stone chippings and sand, which help to give it body. After the ingredients have been mixed with water they are poured into wooden moulds and allowed to set hard. Reinforced concrete is made by allowing concrete to set around steel rods or mesh to give it greater tensile strength, which is required for the construction of large bridges (Figure 13.6) and tall buildings.

Figure 13.6 This bridge in Brazil uses reinforced concrete.

Indirect uses of limestone

Lime manufacture

When calcium carbonate is heated strongly it **thermally dissociates** (breaks up reversibly) to form calcium oxide (lime) and carbon dioxide.

calcium carbonate ⇌ calcium oxide + carbon dioxide

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$

This reaction can go in either direction, depending on the temperature and pressure used. This reaction is an important industrial process and takes place in a lime kiln (Figure 13.7). In excess of 60 million tonnes of calcium oxide are produced worldwide every year.

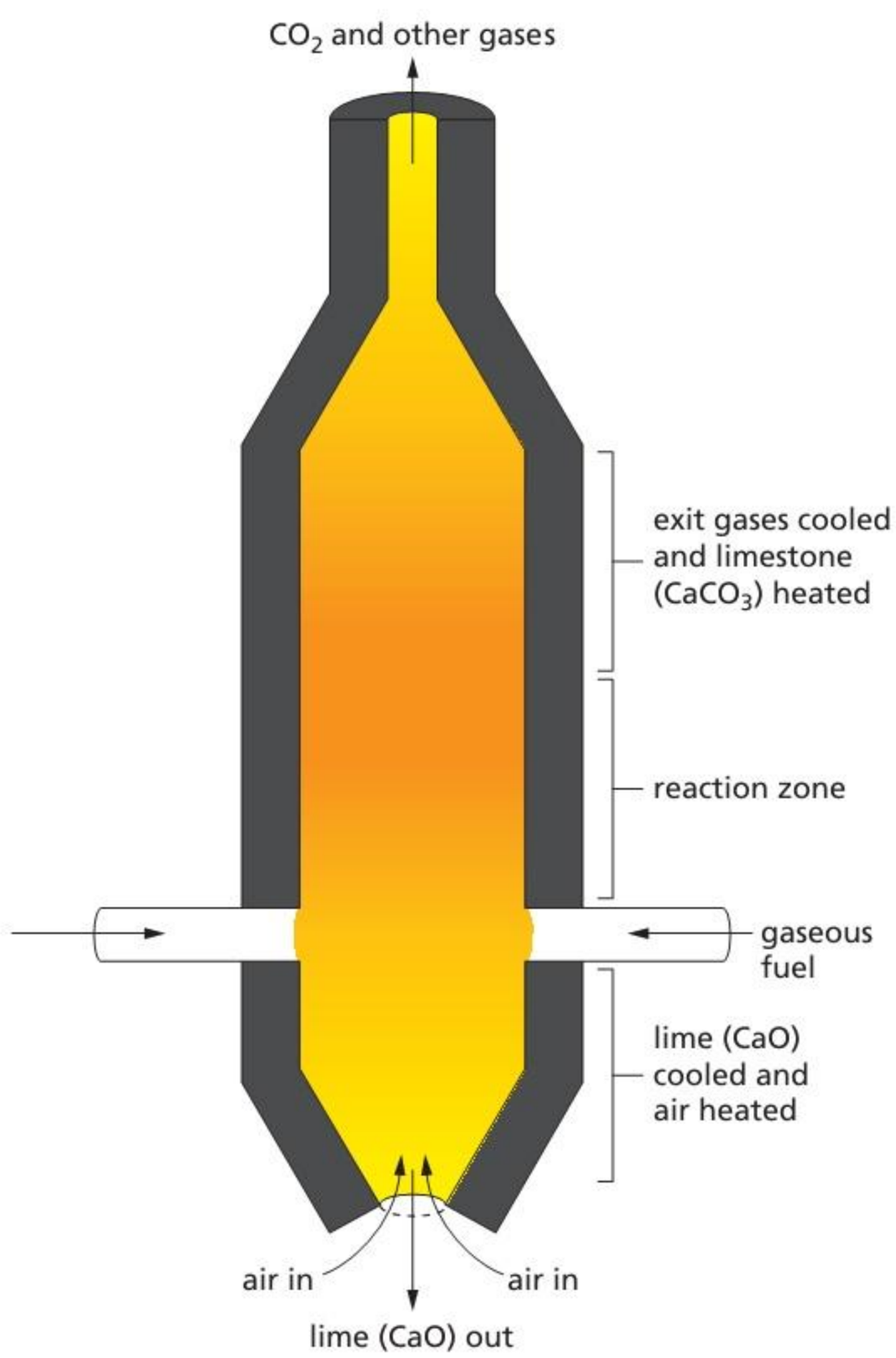


Figure 13.7 The calcium oxide produced from this process is known as quicklime or lime and is used in large quantities in the manufacture of soda glass.

Calcium oxide (CaO) is a base and is still used by some farmers to spread on fields to neutralise soil acidity and to improve drainage of water through soils that contain large amounts of clay. It is also used to neutralise industrial waste products, for example in flue gas desulfurisation. It also has uses as a drying agent in industry and in the manufacture of mouthwash. Soda glass is made by heating sand with soda (sodium carbonate, Na_2CO_3) and lime (Figure 13.8). For further discussion of glasses see Chapter 3, p. 54.

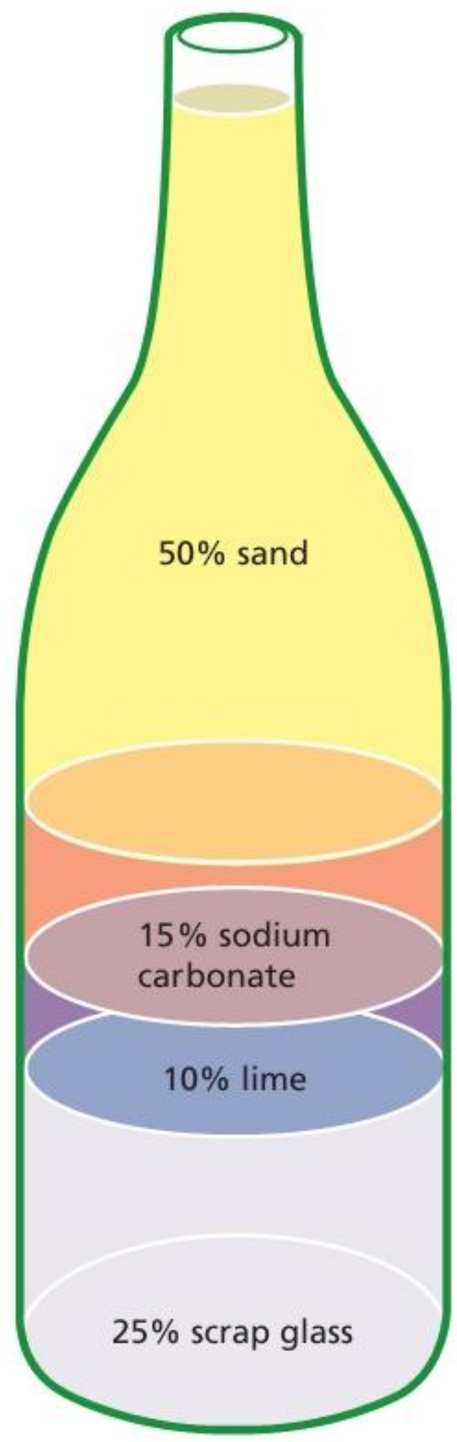


Figure 13.8 The composition of glass.

Large amounts of calcium oxide are also converted into calcium hydroxide ($Ca(OH)_2$) which is called **slaked lime**.

Manufacture of calcium hydroxide – slaked lime

Calcium hydroxide is a cheap industrial alkali (Figure 13.9). It is used in large quantities to make bleaching powder, by some farmers to reduce soil acidity, for neutralising acidic industrial waste products, in the manufacture of whitewash, in glass manufacture and in water purification. Calcium hydroxide, in its white powder form, is produced by adding an equal amount of water to calcium oxide in a carefully controlled reaction. The control is needed because it is a very exothermic reaction.

Figure 13.9 Large-scale production of calcium hydroxide.

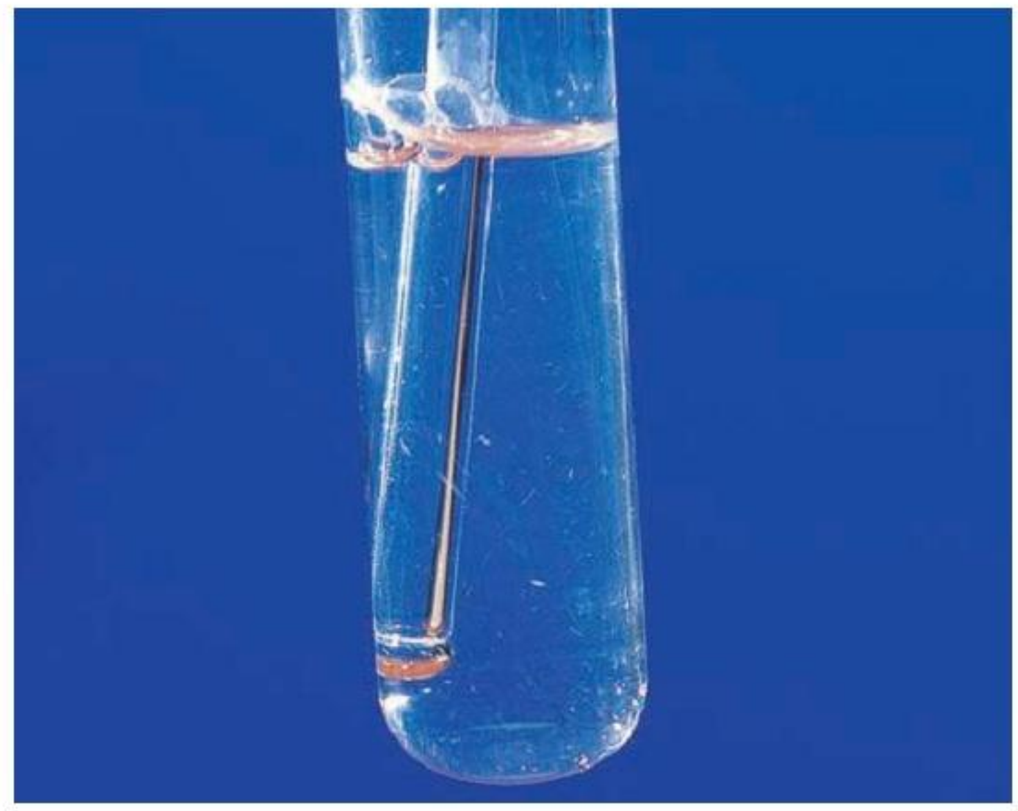
This process can be shown on the small scale in the laboratory by heating a lump of limestone very strongly to convert it to calcium oxide. Water can then be carefully added dropwise to the calcium oxide. An exothermic reaction takes place as the water and calcium oxide react together in this slaking process to form calcium hydroxide.

$$\text{calcium oxide} + \text{water} \rightarrow \text{calcium hydroxide}$$
$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$$

A weak solution of calcium hydroxide in water is called **limewater**. It is used to test for carbon dioxide gas, as a white solid of calcium carbonate is formed if carbon dioxide gas is mixed with it (Figure 13.10):

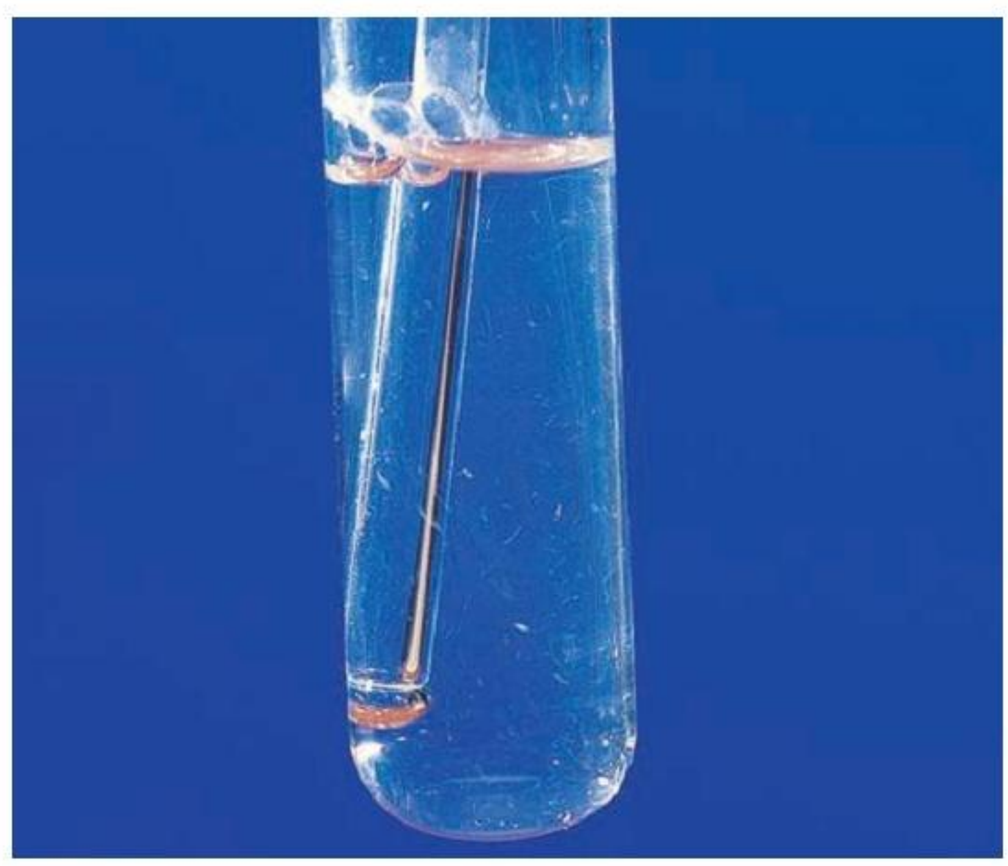
$$\text{calcium hydroxide} + \text{carbon dioxide} \rightarrow \text{calcium carbonate} + \text{water}$$
$$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$

This brings us full cycle from limestone (calcium carbonate, $CaCO_3$), through calcium oxide (CaO) and back to calcium carbonate (limestone). This is known as the 'limestone cycle' (Figure 13.11).



a At the start of the test.

b The limewater turns cloudy.



c If the CO_2 is bubbled for longer the white precipitate dissolves.
Figure 13.10 Testing for carbon dioxide.

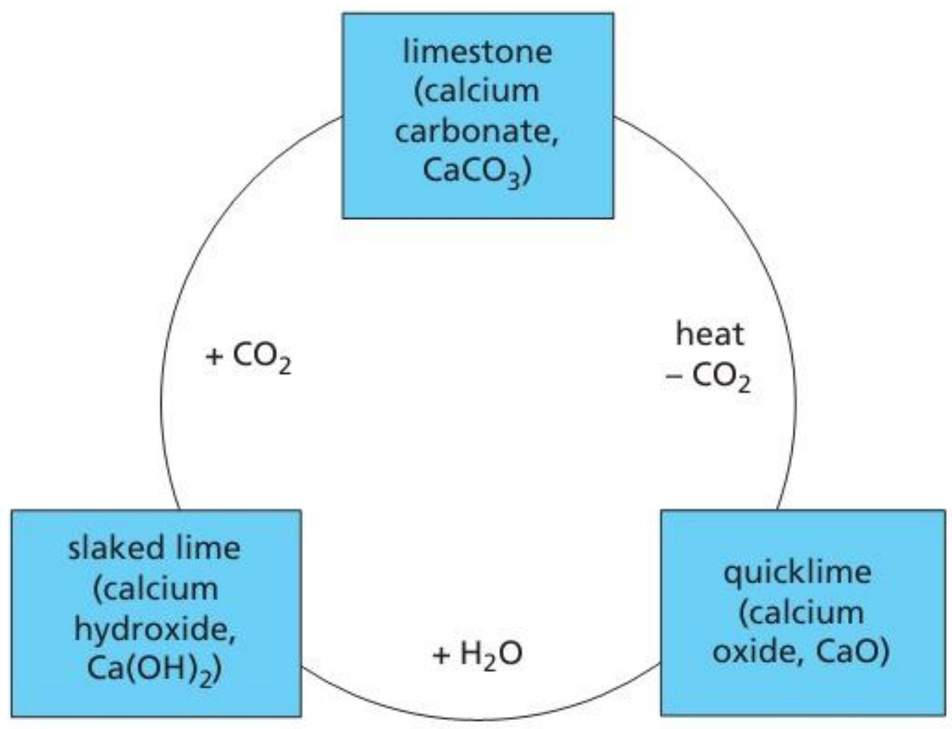


Figure 13.11 The limestone cycle.

If carbon dioxide is bubbled for a further length of time then the white precipitate of calcium carbonate dissolves and a solution of calcium hydrogencarbonate is produced:

calcium carbonate	+	carbon dioxide	+	water	→	calcium hydrogen-carbonate
$CaCO_3(s)$	+	$CO_2(g)$	+	$H_2O(l)$	→	$Ca(HCO_3)_2(aq)$

Calcium hydroxide (slaked lime) is mixed with sand to give mortar. When it is mixed with water and allowed to set a strongly bonded material is formed, which is used to hold bricks together. The hardening of mortar takes place as the following reaction occurs.

calcium hydroxide	+	carbon dioxide	→	calcium carbonate	+	water
$Ca(OH)_2(aq)$	+	$CO_2(g)$	→	$CaCO_3(s)$	+	$H_2O(l)$

Questions

1. Some suggested building alternatives to limestone are sandstone and granite. What would be the benefits of using either of these two materials instead of limestone? For further information, consult the Internet.
2. Devise an experiment that you could carry out in the laboratory to establish whether or not powdered limestone is better at curing soil acidity than calcium oxide.
3. Why is the metallic calcium oxide described as a 'base'?

Carbonates

Carbonates form an important range of compounds. They are all salts of carbonic acid (H_2CO_3) and contain the carbonate ion (CO_3^{2-}). Many of them occur naturally in rock formations. For example, in addition to the calcium carbonate in limestone, chalk and marble, malachite is copper(II) carbonate and dolomite is magnesium carbonate (Figure 13.12).

Figure 13.12 Three naturally occurring carbonates – dolomite with magnesite in front (left) and malachite (right).

Properties of carbonates

- Most metal carbonates thermally decompose when heated to form the metal oxide and carbon dioxide gas (see Chapter 10, p. 152). For example,

copper(II) carbonate	→	copper(II) oxide	+	carbon dioxide
$CuCO_3(s)$	→	$CuO(s)$	+	$CO_2(g)$

  Group I metal carbonates, except for lithium carbonate, do not dissociate on heating. It is generally found that the carbonates of the more reactive metals are much more difficult to dissociate than, for example, copper(II) carbonate. For further discussion of the effect of heat on carbonates see Chapter 10, p. 152.
- Carbonates are generally insoluble in water except for those of sodium, potassium and ammonium.
- Carbonates react with acids to form salts, carbon dioxide and water (Chapter 8, p. 123). For example, calcium carbonate reacts with dilute hydrochloric acid to form calcium chloride, carbon dioxide and water.

calcium carbonate	+	hydrochloric acid	→	calcium chloride	+	carbon dioxide	+	water
$CaCO_3(s)$	+	$2HCl(aq)$	→	$CaCl_2(aq)$	+	$CO_2(g)$	+	$H_2O(l)$

- This reaction is used in the laboratory preparation of carbon dioxide (p. 214). It is also used as a test for a carbonate because the reaction produces carbon dioxide which causes effervescence and if bubbled through limewater turns it chalky white.

Questions

1 Write word and balanced chemical equations to show the effect of heat on cobalt carbonate ($CoCO_3$) and nickel carbonate ($NiCO_3$).

2 Write word and balanced chemical equations for the reaction of dilute hydrochloric acid with:
a copper(II) carbonate b sodium carbonate.

Carbon dioxide

Carbon forms two oxides – carbon monoxide (CO) and carbon dioxide (CO_2). Carbon dioxide is the more important of the two, and in industry large amounts of carbon dioxide are obtained from the liquefaction of air. Air contains approximately 0.03% by volume of carbon dioxide. This value has remained almost constant for a long period of time and is maintained via the **carbon cycle** (Figure 13.13). However, scientists have recently detected an increase in the amount of carbon dioxide in the atmosphere to approximately 0.04%.

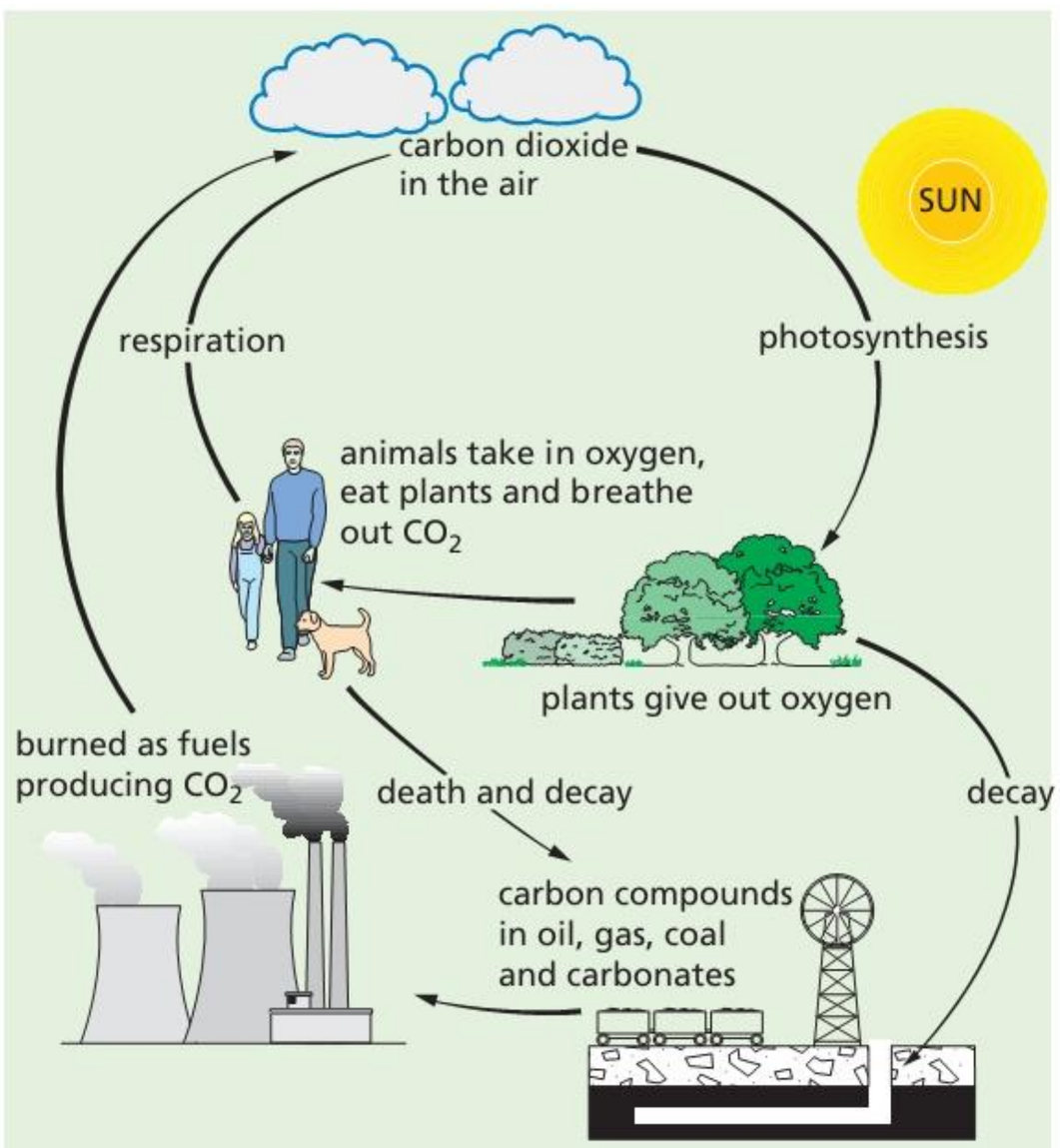


Figure 13.13 The carbon cycle.

Carbon dioxide is produced by burning fossil fuels. It is also produced by all living organisms through **aerobic respiration**. Animals take in oxygen and breathe out carbon dioxide.

glucose + oxygen → carbon dioxide + water + energy

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

Carbon dioxide is taken in by plants through their leaves and used together with water, taken in through their roots, to synthesise sugars. This is the process of **photosynthesis**, and it takes place only in sunlight and only in green leaves, as they contain **chlorophyll** (the green pigment) which catalyses the process.

carbon dioxide + water $\xrightarrow[\text{chlorophyll}]{\text{sunlight}}$ glucose + oxygen

$$6CO_2(g) + 6H_2O(l) \xrightarrow{\text{sunlight, chlorophyll}} C_6H_{12}O_6(aq) + 6O_2(g)$$

The carbon cycle has continued in this manner for millions of years. However, scientists have detected an imbalance in the carbon cycle due to the increase in the amount of carbon dioxide produced through burning fossil fuels and the deforestation of large areas of tropical rain forest. The Earth's climate is affected by the levels of carbon dioxide (and water vapour) in the atmosphere. If the amount of carbon dioxide, in particular, builds up in the air, it is thought that the average temperature of the Earth will rise. This effect is known as the **greenhouse effect** (Figure 13.14).

Some energy from the Sun is absorbed by the Earth and its atmosphere. The remainder is reflected back into space. The energy that is absorbed helps to heat up the Earth. The Earth radiates some heat energy back into space but the 'greenhouse gases', including carbon dioxide, prevent it from escaping. This effect is similar to that observed in a greenhouse where sunlight (visible/ultraviolet radiation) enters through the glass panes but heat (infrared radiation) has difficulty escaping through the glass. Other gases also contribute to the greenhouse effect. One of these is methane, which is produced from agriculture and released from landfill sites as well as rice fields.

The long-term effect of the higher temperatures of the greenhouse effect and the subsequent **global warming** will be the continued gradual melting of ice caps and consequent flooding in low-lying areas of the Earth. There will also be further changes in the weather patterns which will further affect agriculture worldwide.

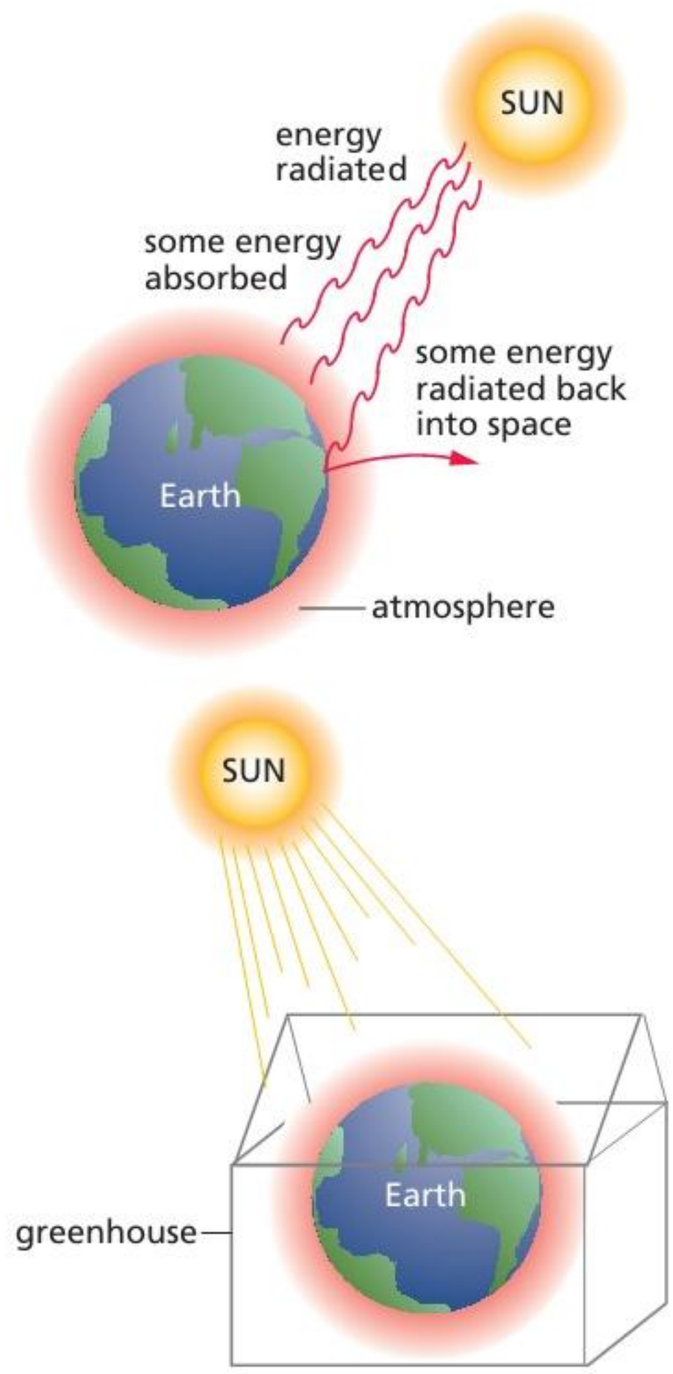


Figure 13.14 The greenhouse effect.

These problems have been recognised by nations worldwide. Recent agreements under the Kyoto Protocol between nations mean that there will be some reduction in the amount of carbon dioxide (and other greenhouse gases) produced over the next few years. However, there is still a long way to go.

Questions

1 'Burning fossil fuel in an aircraft engine is significantly worse than burning the fuel at ground level.'
Discuss this statement in terms of the greenhouse effect and global warming.

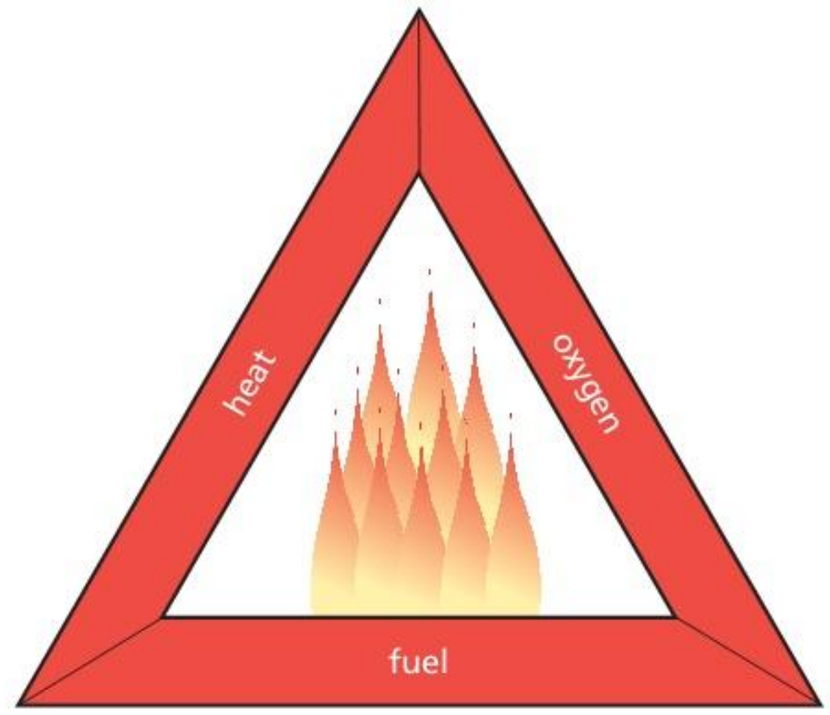
2 Use your research skills to find out about:
   a the carbon trading scheme
   b low-carbon homes
   c the Kyoto Protocol.

3 Use your research skills to discover changes that are taking place to the ice caps and weather patterns worldwide.

Uses of carbon dioxide

Carbon dioxide has some important uses.

- **Carbonated drinks.** Large quantities are used to make soda and mineral waters. The carbon dioxide gas is bubbled into the liquid under pressure, which increases its solubility.
- **Fire extinguishers.** It is used in extinguishers for use on electrical fires. Carbon dioxide is denser than air and forms a layer around the burning material. It covers the fire and starves it of oxygen. Carbon dioxide does not burn and so the fire is put out (Figure 13.15).



a Any fire needs fuel, oxygen and heat. If any of these is removed, the fire triangle is destroyed and the fire will be extinguished.

b Carbon dioxide fire extinguishers starve the fire of oxygen. They are used mainly for electrical fires.
Figure 13.15

- **Refrigerants.** Solid carbon dioxide (dry ice) is used for refrigerating ice cream, meat and soft fruits. It is used for this purpose because it is colder than ice and it sublimes (p. 5), and so it does not pass through a potentially damaging liquid stage.
- **Special effects.** Carbon dioxide is used to create the 'smoke' effect you may see at pop concerts and on television. Dry ice is placed in boiling water and it forms thick clouds of white 'smoke' (Figure 1.9, p. 5). It stays close to the floor due to the fact that carbon dioxide is denser than air.
- **Heat transfer agents.** Carbon dioxide gas is used for transferring heat in some nuclear power stations.

Laboratory preparation of carbon dioxide gas

In the laboratory the gas is made by pouring dilute hydrochloric acid on to marble chips ($CaCO_3$).

calcium carbonate	+	hydrochloric acid	→	calcium chloride	+	water	+	carbon dioxide
$CaCO_3(s)$	+	$2HCl(aq)$	→	$CaCl_2(aq)$	+	$H_2O(l)$	+	$CO_2(g)$

A suitable apparatus for preparing carbon dioxide is shown in Figure 13.16.

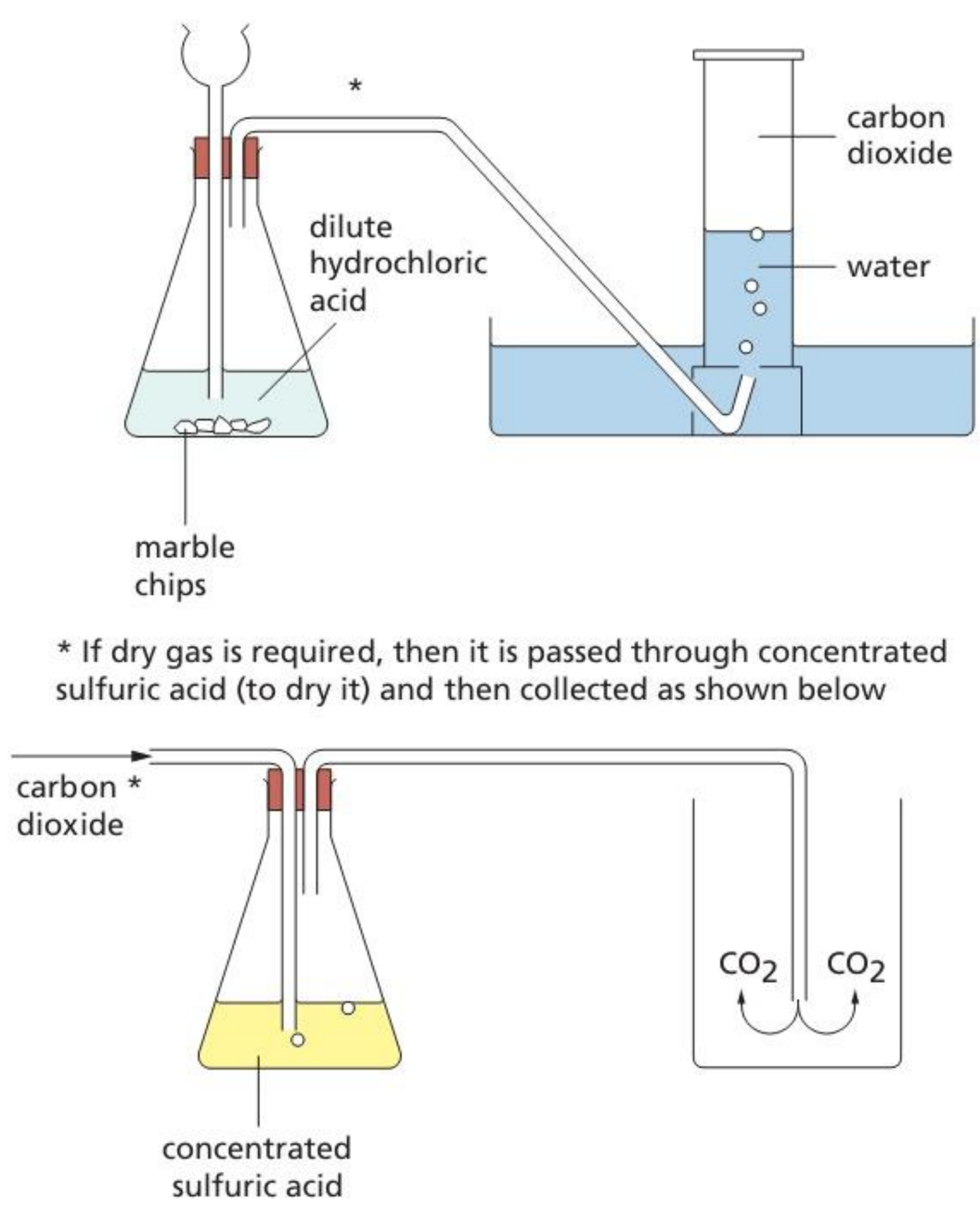


Figure 13.16 Preparation and collection of carbon dioxide gas.

Question

1 List the important uses of carbon dioxide and for each use you have given in your answer explain why carbon dioxide is used.

Properties of carbon dioxide gas

Physical properties

Carbon dioxide is:

- a colourless gas
- sparingly soluble in water
- denser than air.

Chemical properties

- When bubbled into water it dissolves slightly and some of the carbon dioxide reacts, forming a solution of the weak acid carbonic acid which shows a pH of 4 or 5.

water	+	carbon dioxide	⇌	carbonic acid
$H_2O(l)$	+	$CO_2(g)$	⇌	$H_2CO_3(aq)$

- It will support the combustion only of strongly burning substances such as magnesium. This burning reactive metal decomposes the carbon dioxide to provide oxygen for its continued burning in the gas. This reaction is accompanied by much crackling (Figure 13.17).

Figure 13.17 When magnesium burns in carbon dioxide gas, magnesium oxide (white) and carbon (black) are produced.

magnesium	+	carbon dioxide	→	magnesium oxide	+	carbon
$2Mg(s)$	+	$CO_2(g)$	→	$2MgO(s)$	+	$C(s)$

This reaction is a good example of a **redox** process (Chapter 2, p. 14). Which of these substances has been oxidised and which has been reduced? Which of these substances is the oxidising agent and which is the reducing agent?

- When carbon dioxide is bubbled through limewater (calcium hydroxide solution), a white precipitate is formed. This white solid is calcium carbonate ($CaCO_3$). This reaction is used as a test to show that a gas is carbon dioxide.

carbon dioxide	+	calcium hydroxide	→	calcium carbonate	+	water
$CO_2(g)$	+	$Ca(OH)_2(aq)$	→	$CaCO_3(s)$	+	$H_2O(l)$

If carbon dioxide is bubbled through this solution continuously then it will eventually become clear. This is because of the formation of soluble calcium hydrogencarbonate solution.

calcium carbonate	+	water	+	carbon dioxide	→	calcium hydrogen-carbonate
$CaCO_3(s)$	+	$H_2O(l)$	+	$CO_2(g)$	→	$Ca(HCO_3)_2(aq)$

- Carbon dioxide reacts with strong alkalis, such as sodium hydroxide, to form carbonates. A solution of sodium hydroxide can be used to absorb carbon dioxide from the air. If excess carbon dioxide is bubbled through a solution of an alkali then a white precipitate may be obtained. This is due to the formation and precipitation of the sodium hydrogencarbonate:

sodium carbonate	+	water	+	carbon dioxide	→	sodium hydrogen-carbonate
$Na_2CO_3(aq)$	+	$H_2O(l)$	+	$CO_2(g)$	→	$2Na(HCO_3)_2(s)$

Sodium hydrogencarbonate is quite a useful substance. Because it is thermally unstable (easily decomposed by heat) it is used as a 'raising agent' in baking (Chapter 8, p. 126).

Questions

1. Describe an experiment that you could carry out to show that carbonic acid is a weak acid.
2. Why is it not possible to use dilute sulfuric acid to make carbon dioxide from limestone in the laboratory?
3. When carbon dioxide is 'poured' from a gas jar on to a burning candle the candle goes out. What properties of carbon dioxide does this experiment show?

Checklist

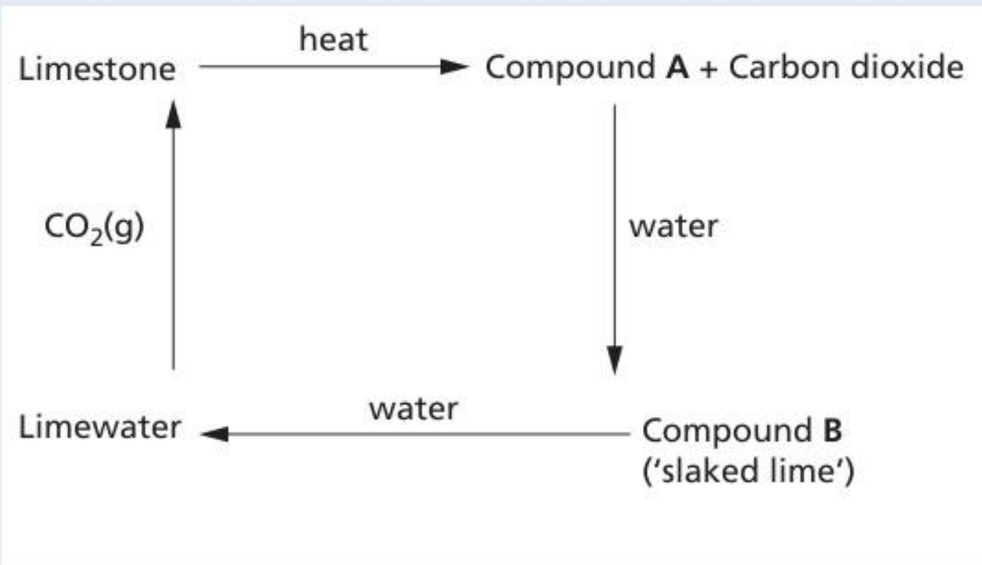
After studying Chapter 13 you should know and understand the following terms.

- **Aerobic respiration** This is respiration that takes place in air.
- **Carbonate** A salt of carbonic acid containing the carbonate ion, CO_3^{2-}, for example $CuCO_3$.
- **Carbon cycle** The complex series of processes through which all of the carbon atoms in existence rotate.
- **Chlorophyll** The green substance in plants that acts as a catalyst in photosynthesis.
- **Global warming** The increased average temperature on Earth caused by the greenhouse effect.
- **Greenhouse effect** The absorption of reflected infrared radiation from the Earth by gases in the atmosphere such as carbon dioxide (a greenhouse gas) leading to atmospheric or global warming.
- **Lime** A white solid known chemically as calcium oxide (CaO). It is produced by heating limestone. It is used to counteract soil acidity and to manufacture calcium hydroxide (slaked lime). It is also used as a drying agent in industry.
- **Photosynthesis** The opposite of respiration. The process by which plants synthesise sugars.
- **Raw materials** The basic materials from which a product is made. For example, the raw materials for the Haber process are nitrogen and hydrogen.
- **Thermal decomposition** The breakdown of a substance under the influence of heat.
- **Thermal dissociation** The reversible breaking up of a substance under the influence of heat.

Inorganic carbon chemistry

Additional questions

1 This question is about the limestone cycle.



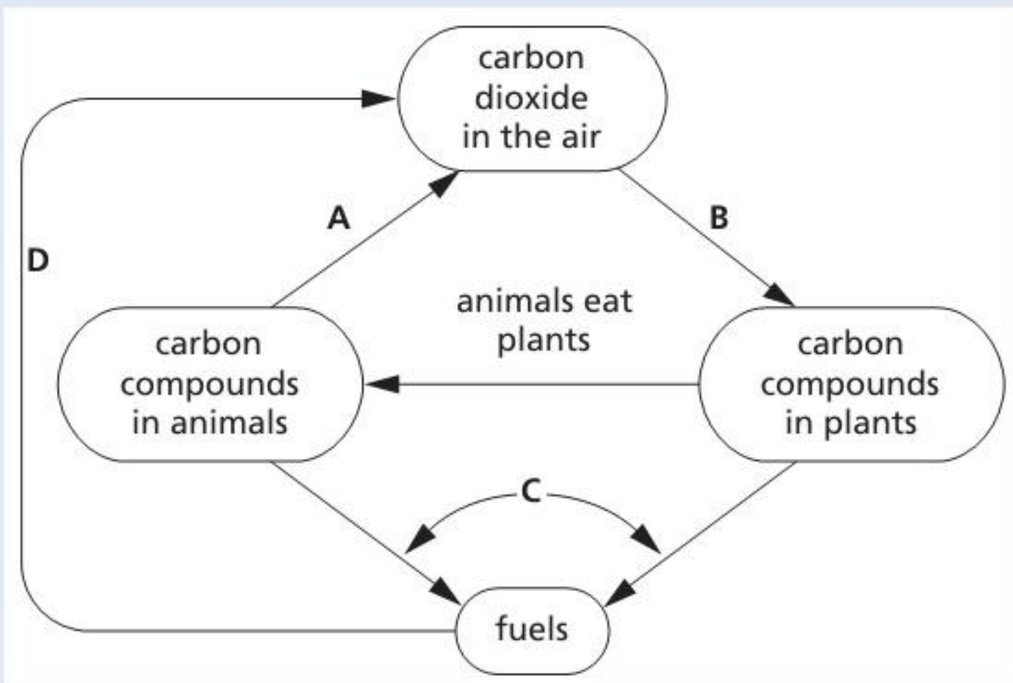
a Name and give the formula of:
(i) compound **A** **(ii)** compound **B**.

b Write balanced chemical equations for the formation of both compounds **A** and **B**.

c Name and give the symbol/formula for the ions present in limewater.

d Describe with the aid of a balanced chemical equation what happens if carbon dioxide is bubbled through limewater until there is no further change.

2 The diagram shown below is a simplified version of the carbon cycle.



a Name the processes **A**, **B**, **C** and **D**.

b Write balanced chemical equations to represent the processes taking place in **A** and **B**.

c There are a number of fuels which could be placed in the 'fuels' box. Name three such fuels.

d 'The air contains approximately 0.03% of carbon dioxide and this amount is almost constant.' Explain why this is a true statement.

e **(i)** What is the 'greenhouse effect'?
(ii) Discuss the possible consequences of the greenhouse effect if nothing is done to counteract it.

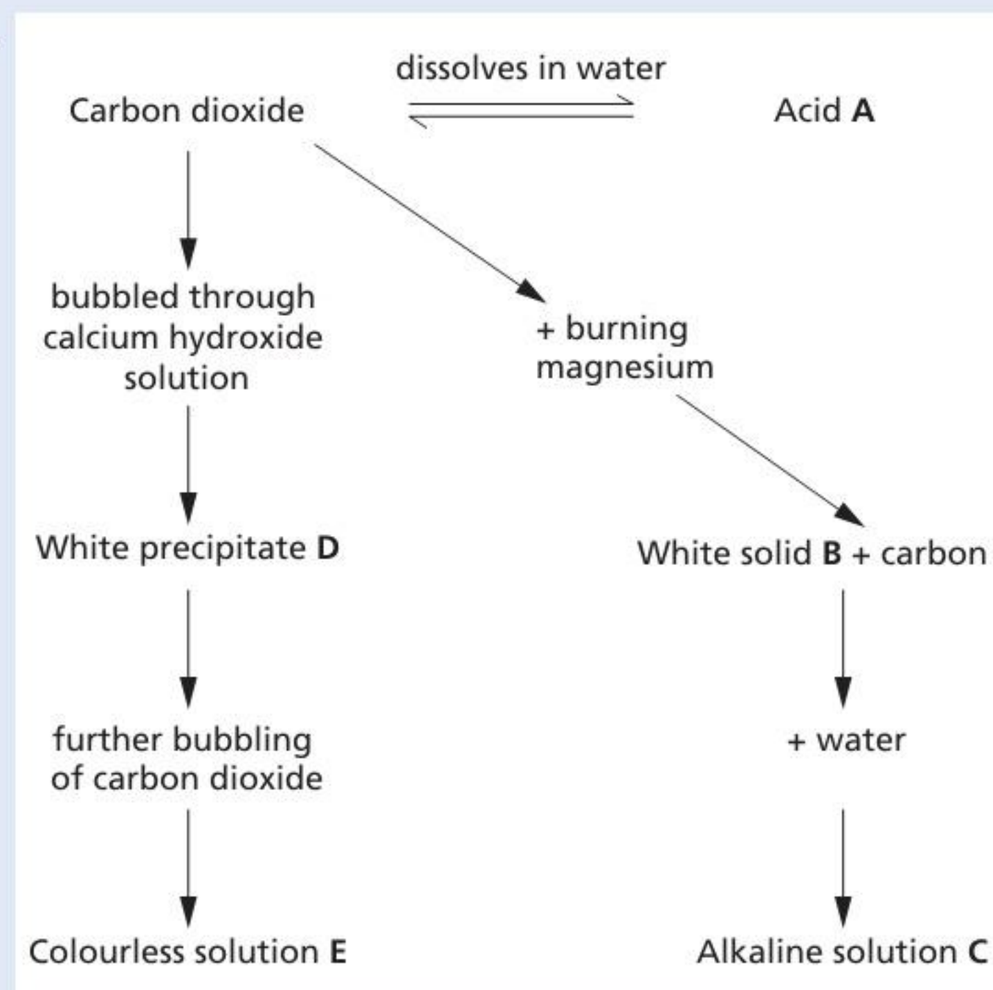
3 Limestone is an important raw material used in many different industries.

a One of the properties of limestone is that it reacts with acids.
(i) Why do farmers spread powdered limestone on their fields?
(ii) How can buildings made of limestone be affected by 'acid rain'?
(iii) Write an ionic equation which would represent the reactions taking place in both **(i)** and **(ii)**.

b Limestone is used in the manufacture of iron (Chapter 10, p. 158).
(i) Why is it added to the blast furnace along with coke and haematite?
(ii) Write chemical equations for the reactions it is involved in, in this process.

c **(i)** Name a building material made by heating a mixture of limestone and clay in a rotary kiln.
(ii) What substance is the dry product produced in the rotary kiln added to?
Explain why this substance is added.

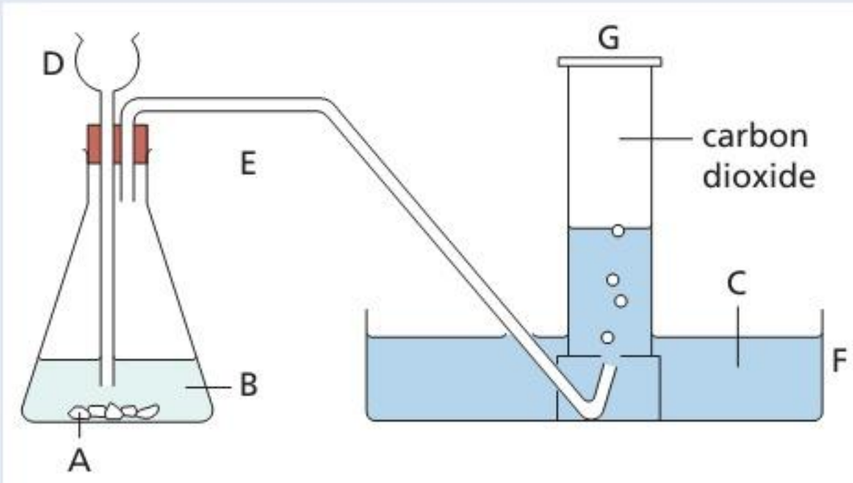
4



a Name and give the formulae of substances **A** to **E**.

b Write balanced chemical equations for the reactions in which compounds **B**, **C** and **E** were formed.
c Where would you expect to find acid **A**?
d Universal indicator solution was added to solution **C**. What colour did it go?
e Upon addition of dilute hydrochloric acid to solution **C**, a neutralisation reaction took place.
(i) Write a balanced chemical equation for the reaction taking place.
(ii) Name the salt produced in this reaction.

5 The following question is about carbon dioxide.



a Name and give the formula of each of the substances **A**, **B** and **C**.
b Identify by name the different pieces of apparatus **D**, **E**, **F** and **G**.
c Draw and label the apparatus that should be used if a dry sample of carbon dioxide gas was required.
d When a gas jar containing carbon dioxide is held over a burning wooden splint and the cover removed, the flame goes out. State two properties of carbon dioxide illustrated by this observation.
e Carbon dioxide is also produced when zinc carbonate is heated strongly.
(i) Write a balanced chemical equation for the reaction taking place.
(ii) Name the process which is taking place as the zinc carbonate is heated.
(iii) Calculate the volume of carbon dioxide that would be produced (measured at room temperature and pressure (rtp)) if 12.5 g of zinc carbonate were heated strongly.
(One mole of any gas occupies 24 dm^3 at rtp.
A_r: C = 12; O = 16; Zn = 65)

6 a Give the names and formulae of the two major greenhouse gases.
b Name a natural source of the gases you have named in part **a**.
c Name a man-made source of the gases you have named in part **a**.

7 Use the words and formulae below to complete the following passage about limestone and its uses.

CaO building $CaCO_3$ acids
thermal CO_2 oxide acidity
quarries quicklime calcium

Limestone is a very common rock. The main chemical in limestone is _____ carbonate (chemical formula, _____). Limestone is obtained from _____ . The uses of limestone are many and varied. For example, because it is easy to cut into blocks it is useful as a _____ material. Also it will react with _____ and so it is used in the powder form to reduce soil _____ . When limestone is heated very strongly, it breaks down into _____ (chemical name, calcium _____) and carbon dioxide (chemical formula, _____). This kind of reaction is known as _____ decomposition.

8 Lime (calcium oxide) is produced in very large quantities in a lime kiln. The equation for the reaction is:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

a How much limestone would be needed to produce 61.60 tonnes of lime?
(A_r: C = 12; O = 16; Ca = 40)
b Why is the carbon dioxide swept out of the lime kiln?
c Give three uses of lime.
d (i) What problems are associated with the large-scale quarrying of limestone?
(ii) What steps have been taken to overcome or reduce the problems you have outlined in **(i)**?

14 Organic chemistry 1

Alkanes
Naming the alkanes
Structural isomerism

The chemical behaviour of alkanes
The ozone hole problem
Methane – another greenhouse gas!
Other uses of alkanes

Alkenes
Naming the alkenes
Where do we get alkenes from?

The chemical behaviour of alkenes
Addition reactions

A special addition reaction of alkene molecules
Other addition polymers
Thermosoftening and thermosetting plastics
Disposal of plastics

Checklist

Additional questions

A lot of the compounds that are present in living things have been found to be compounds containing carbon (Figure 14.1). These are known as **organic compounds**. All living things are made from organic compounds based on chains of carbon atoms which are not only covalently bonded to each other but also covalently bonded to hydrogen, oxygen and/or other elements. The organic compounds are many and varied. Some scientists suggest that there are more than ten million known organic compounds.

Figure 14.1 Living things contain organic compounds.

You saw in Chapter 6 that crude oil is made up of a complex mixture of hydrocarbon compounds. This mixture is a very important raw material in the organic chemical industry. Our world would be a very different place without these substances. Can you imagine life without the substances from oil and natural gas? There would be no oil-based fuels, such as petrol and diesel, and the chemical industry, particularly the plastics industry, would also suffer since substances extracted from oil are used as raw materials for many plastics.

Alkanes

Most of the hydrocarbons in crude oil belong to the family of compounds called **alkanes**. The molecules within the alkane family contain carbon atoms covalently bonded to four other atoms by single bonds (Figure 14.2). Because these molecules possess only single bonds they are said to be **saturated**, as no further atoms can be added. This can be seen in the bonding scheme for methane (Figure 14.3). The physical properties of the first six members of the alkane family are shown in Table 14.1.

You will notice from Figure 14.2 and Table 14.1 that the compounds have a similar structure and similar name endings. They also behave chemically in a similar way. A family with these factors in common is called a **homologous series**.

All the members of a homologous series can also be represented by a general formula. In the case of the alkanes the general formula is:

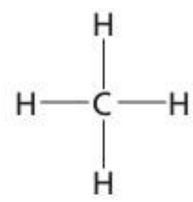
$$C_nH_{2n+2}$$

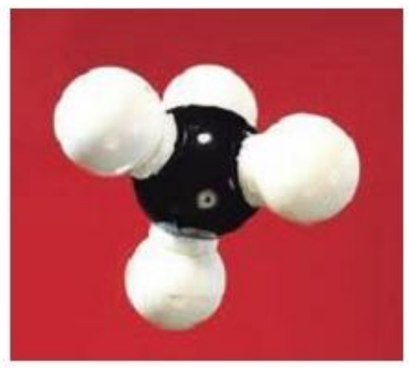
where n is the number of carbon atoms present.

As you go up a homologous series, in order of increasing number of carbon atoms, the physical properties of the compounds gradually change. For example, the melting and boiling points of the alkanes shown in Table 14.1 gradually increase. This is due to an increase in the intermolecular forces (van der Waals' forces) as the size and mass of the molecule increases (Chapter 3, p. 49).

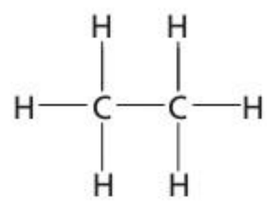
Under normal conditions molecules with up to four carbon atoms are gases, those with between five and 16 carbon atoms are liquids, while those with more than 16 carbon atoms are solids.

methane

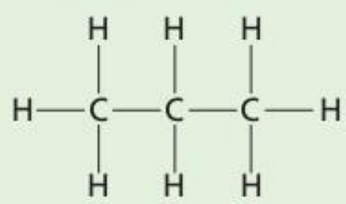




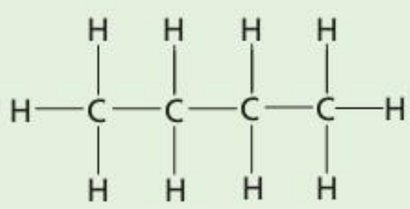
ethane

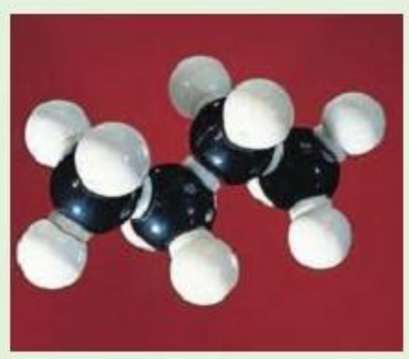


propane

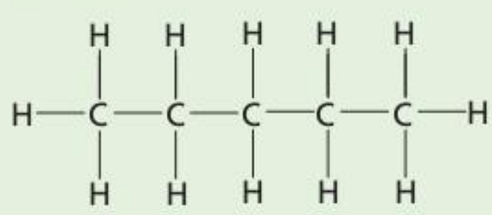


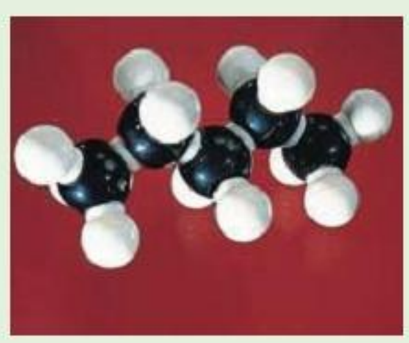
butane



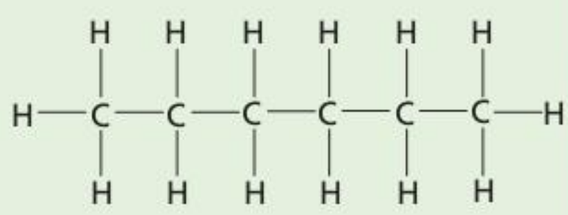


pentane





hexane



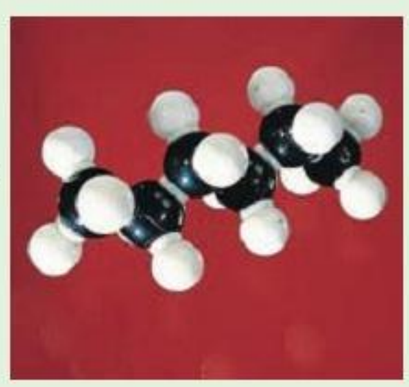
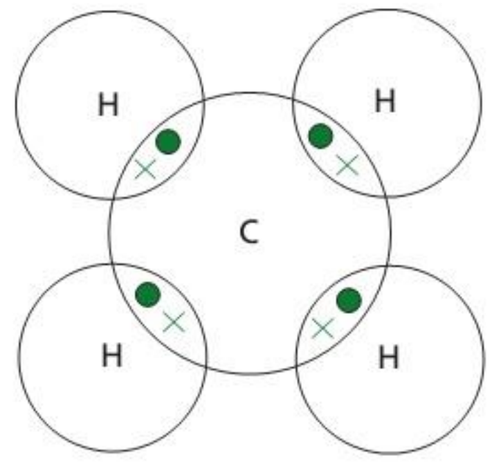


Figure 14.2 The fully displayed (or structural) formulae and molecular models of the first six alkanes.



methane molecule (CH_4)

Figure 14.3 The covalent bonding scheme for methane.

Table 14.1 Some alkanes and their physical properties.

Alkane	Formula	Melting point/ °C	Boiling point/ °C	Physical state at room temperature
Methane	CH_4	−182	−162	Gas
Ethane	C_2H_6	−183	−89	Gas
Propane	C_3H_8	−188	−42	Gas
Butane	C_4H_{10}	−138	0	Gas
Pentane	C_5H_{12}	−130	36	Liquid
Hexane	C_6H_{14}	−95	69	Liquid

Questions

1 Estimate the boiling points for the alkanes with formulae:
   a C_7H_{16} **b** C_8H_{18}.
2 Name the alkanes which have the following formulae:
   a C_7H_{16} **b** $C_{10}H_{22}$.

Naming the alkanes

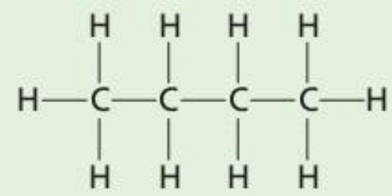
All the alkanes have names ending in *-ane*. The rest of the name tells you the number of carbon atoms present in the molecule. For example, the compound whose name begins with:

- *meth-* has one carbon atom
- *eth-* has two carbon atoms
- *prop-* has three carbon atoms
- *but-* has four carbon atoms
- *pent-* has five carbon atoms

and so on.

Structural isomerism

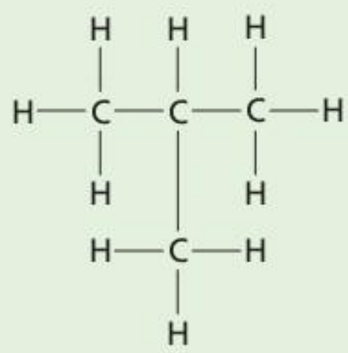
Sometimes it is possible to write more than one structural formula to represent a molecular formula. The structural formula of a compound shows how the atoms are joined together by the covalent bonds. For example, there are two different compounds with the molecular formula C_4H_{10}. The structural formulae of these two substances along with their names and physical properties are shown in Figure 14.4.



melting point −138°C

boiling point 0°C

a butane



melting point −159°C

boiling point −12°C

b 2-methylpropane

Figure 14.4 The isomers of C_4H_{10}.

Compounds such as those in Figure 14.4 are known as **isomers**. Isomers are substances which have the same molecular formula but different structural formulae. The different structures of the compounds shown in Figure 14.4 have different melting and boiling points. Molecule **b** contains a branched chain and has a lower melting point than molecule **a**, which has no branched chain. All the alkane molecules with four or more carbon atoms possess isomers. Perhaps now you can see why there are so many different organic compounds!

Question

1 Draw the structural formulae for the isomers of C_5H_{12}.

The chemical behaviour of alkanes

Alkanes are rather unreactive compounds. For example, they are generally not affected by alkalis, acids or many other substances. Their most important property is that they burn easily.

Gaseous alkanes, such as methane, will burn in a good supply of air, forming carbon dioxide and water as well as plenty of heat energy.

methane + oxygen → carbon dioxide + water + energy

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

The gaseous alkanes are some of the most useful fuels. Methane, better known as natural gas, is used for cooking as well as for heating offices, schools and homes (Figure 14.5a). Propane and butane burn with very hot flames and they are sold as liquefied petroleum gas (LPG). In rural areas where there is no supply of natural gas, central heating systems can be run on propane gas (Figure 14.5b). Butane, sometimes mixed with propane, is used in portable blowlamps and in gas lighters.

a This is burning methane.

b Central heating systems can be run on propane.

Figure 14.5

Another useful reaction worth noting is that between the alkanes and the halogens. For example, methane and chlorine react in the presence of sunlight (or ultraviolet light). The ultraviolet light splits the chlorine molecules into atoms. When this type of reaction takes place, these atoms are called **free radicals** and they are very reactive.

chlorine gas $\xrightarrow{\text{sunlight}}$ chlorine atoms (free radicals)

$$Cl_2(g) \longrightarrow 2Cl(g)$$

The chlorine atoms then react further with methane molecules, and a hydrogen chloride molecule is produced along with a methyl free radical.

chlorine atom + methane → methyl radical + hydrogen chloride

$Cl(g) + CH_4(g) \rightarrow CH_3(g) + HCl(g)$

The methyl free radical reacts further.

methyl radical + chlorine gas → chloromethane + chlorine atom

$CH_3(g) + Cl_2(g) \rightarrow CH_3Cl(g) + Cl(g)$

This chlorine free radical, in turn, reacts further and the process continues until all the chlorine and the methane have been used up. This type of process is known as a **chain reaction** and it is very fast. The overall chemical equation for this process is:

methane + chlorine → chloromethane + hydrogen chloride

$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$

We can see that one hydrogen atom of the methane molecule is **substituted** by a chlorine atom. This type of reaction is known as a **substitution reaction**.

Because we cannot control the chlorine free radicals produced in this reaction, we also obtain small amounts of other 'substituted' products – CH_2Cl_2 (dichloromethane), $CHCl_3$ (trichloromethane or chloroform) and CCl_4 (tetrachloromethane) – by further reactions such as those shown below.

chloromethane + chlorine atom → chloromethyl radical + hydrogen chloride

$CH_3Cl(g) + Cl(g) \rightarrow CH_2Cl(g) + HCl(g)$

chloromethyl radical + chlorine gas → dichloromethane + chlorine atom

$CH_2Cl(g) + Cl_2(g) \rightarrow CH_2Cl_2(g) + Cl(g)$

Many of these so-called **halogenoalkanes** are used as solvents. For example, dichloromethane is used as a solvent in paint stripper (Figure 14.6).

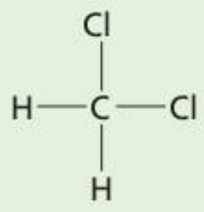


Figure 14.6 Dichloromethane is used as a solvent in paint stripper.

Early anaesthetics relied upon trichloromethane, $CHCl_3$, or chloroform. Unfortunately, this anaesthetic had a severe problem since the lethal dose was only slightly higher than that required to anaesthetise the patient. In 1956, halothane was discovered by chemists working at ICI. This is a compound containing chlorine, bromine and fluorine. Its formula is $CF_3CHBrCl$. However, even this is not the perfect anaesthetic since evidence suggests that prolonged exposure to this substance may cause liver damage. The search continues for even better anaesthetics.

A group of compounds were discovered in the 1930s and were called the chlorofluorocarbons (CFCs). Because of their inertness they found many uses, especially as a propellant in aerosol cans. CFC-12 or dichlorodifluoromethane, CF_2Cl_2, was one of the most popular CFCs in use in aerosols. Scientists believe that CFCs released from aerosols are destroying the ozone layer and steps are being taken to reduce this threat.

The ozone hole problem

Our atmosphere protects us from harmful ultraviolet radiation from the Sun. This damaging radiation is absorbed by the relatively thin ozone layer found in the stratosphere (Figure 14.7).

Large holes have recently been discovered in the ozone layer over Antarctica, Australasia and Europe (Figure 11.5, p. 173). Scientists think that these holes have been produced by CFCs such as

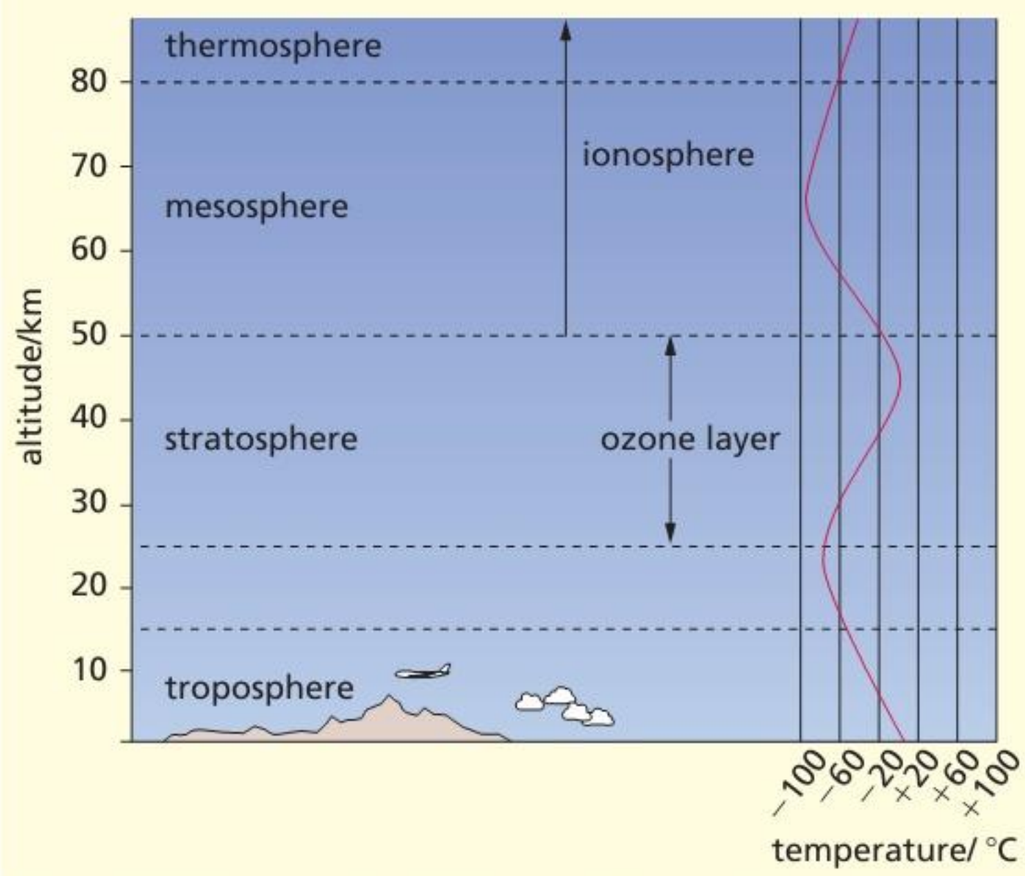


Figure 14.7 The ozone layer is between 25 km and 50 km above sea level.

CFC-12. CFCs escape into the atmosphere and, because of their inertness, remain without further reaction until they reach the stratosphere and the ozone layer. In the stratosphere the high-energy ultraviolet radiation causes a chlorine atom to split off from the CFC molecule. This chlorine atom, or free radical, then reacts with the ozone.

$$Cl(g) + O_3(g) \rightarrow OCl(g) + O_2(g)$$

This is not the only problem with CFCs. They are also significant 'greenhouse gases' (Chapter 13, p. 212). The 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 intended agreements are adhered to then the ozone layer will recover by 2050.

Methane – another greenhouse gas!

Methane, the first member of the alkanes, occurs naturally. Cows produce it in huge quantities when digesting their food. It is also formed by growing rice. Like carbon dioxide, it is a greenhouse gas (Chapter 13, p. 212) because it acts like the glass in a greenhouse – it will let in heat from the Sun but will not let all of the heat back out again. It is thought that the greenhouse effect may contribute to climate change, which could have disastrous effects for life on this planet.

Question

1 Use your research skills to find out:
   a any other sources of methane found in nature
   b how climate change might affect your particular environment.

Methane – the 21st century chemical

Methane is finding another use, which sets it apart from the rest of the alkanes, in the production of nanotubes for power generation and futuristic circuitry. A common way to grow nanotubes is from a mixture of methane gas and iron nanoparticles. When heated to around 700–800°C in a vacuum the iron acts as a catalyst firstly removing carbon from the methane. When the surface of the iron becomes saturated, nanotubes begin to be formed. If He or Ar is added to the vacuum chamber then 'metallic' (electrical conducting) nanotubes are made in the majority; this accounts for up to 91% of the nanotubes produced (Figure 14.8).

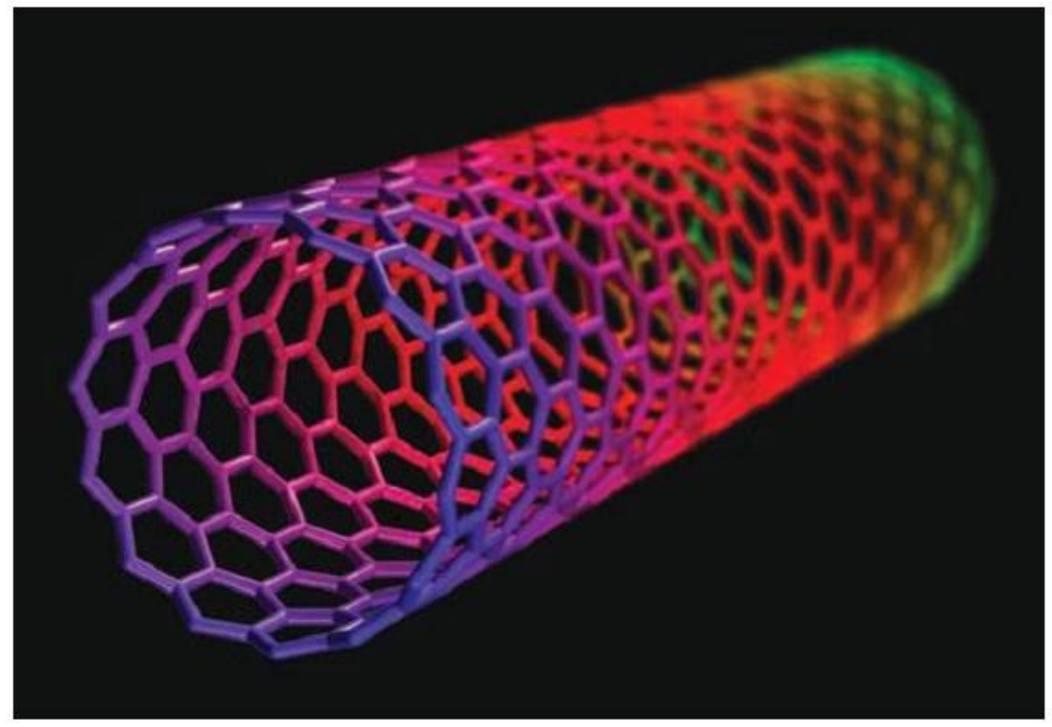


Figure 14.8 These nanotubes are being used in electrical circuitry.

Other uses of alkanes

Besides their major use as fuels (p. 90), some of the heavier alkanes are used as waxes in candles, as lubricating oils and in the manufacture of another family of hydrocarbons – the alkenes.

Questions

1 Write a balanced chemical equation to represent the combustion of propane.
2 In what mole proportions should chlorine and methane be mixed to produce:
   a mainly chloromethane?
   b mainly tetrachloromethane?
3 Describe a method you would use to separate chloromethane from the other possible reaction products when methane reacts with chlorine.
4 Explain why it is the C—Cl bond and not the C—F bond in CFCs that is disrupted in the stratosphere.

Alkenes

Alkenes form another homologous series of hydrocarbons of the general formula C_nH_{2n} where n is the number of carbon atoms. The alkenes are more reactive than the alkanes because they each contain a double covalent bond between the carbon atoms (Figure 14.9). Molecules that possess a double covalent bond of this kind are said to be

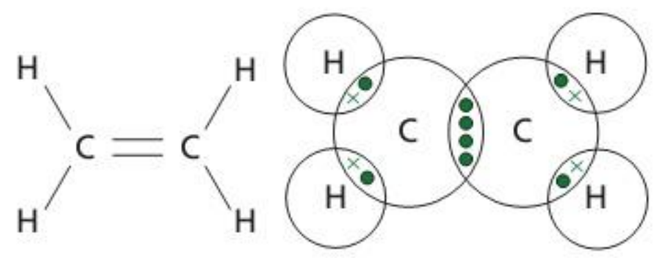


Figure 14.9 The bonding in ethene, the simplest alkene.

unsaturated, because it is possible to break one of the two bonds to add extra atoms to the molecule.

The chemical test to show the difference between saturated and unsaturated hydrocarbons is discussed on p. 225.

Naming the alkenes

All alkenes have names ending in *-ene*. Alkenes, especially ethene, are very important industrial chemicals. They are used extensively in the plastics industry and in the production of alcohols such as ethanol and propanol. See Table 14.2 and Figure 14.10.

Table 14.2 The first three alkenes and their physical properties.

Alkene	Formula	Melting point/°C	Boiling point/°C	Physical state at room temperature
Ethene	C_2H_4	−169	−104	Gas
Propene	C_3H_6	−185	−47	Gas
Butene	C_4H_8	−184	−6	Gas

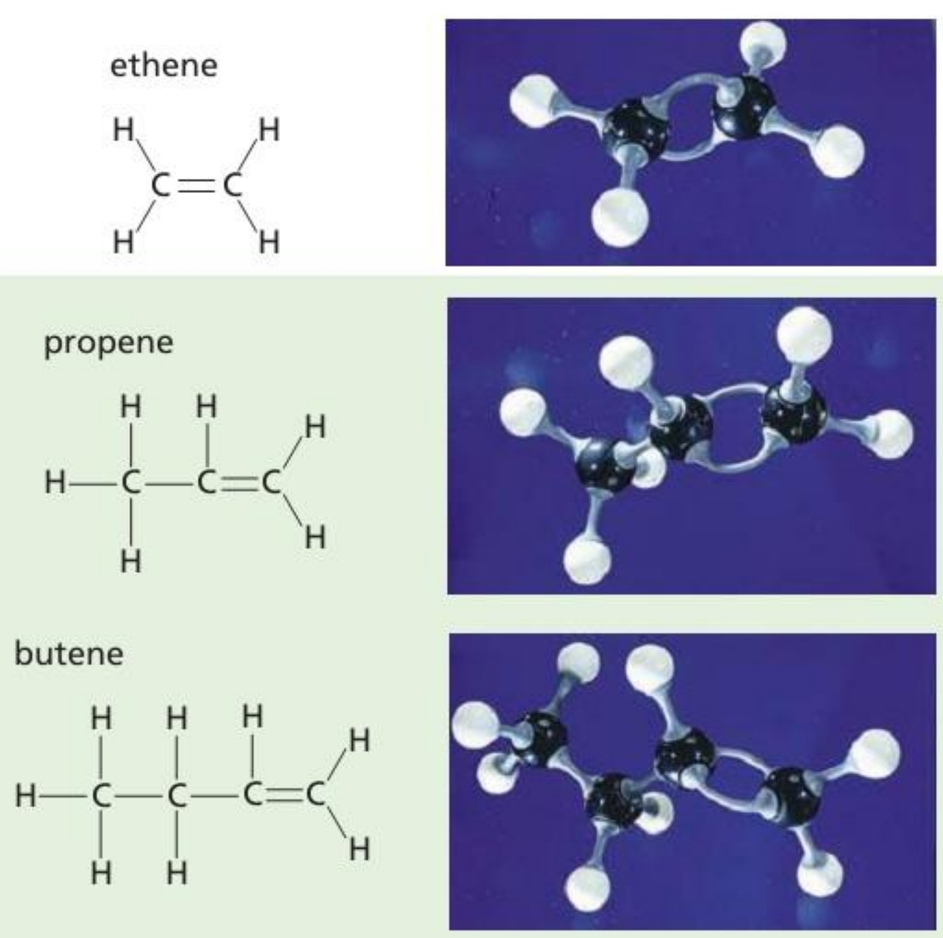


Figure 14.10 Structure and shape of the first three alkenes.

Where do we get alkenes from?

Very few alkenes are found in nature. Most of the alkenes used by the petrochemical industry are obtained by breaking up larger, less useful alkane molecules obtained from the fractional distillation of crude oil. This is usually done by a process called **catalytic cracking**. In this process the alkane molecules to be 'cracked' (split up) are passed over a mixture of aluminium and chromium oxides heated to about 500 °C.

$$\underset{\text{(found in kerosene)}}{\text{dodecane}} \longrightarrow \underset{\text{shorter alkane}}{\text{decane}} + \underset{\text{alkene}}{\text{ethene}}$$

$$C_{12}H_{26}(g) \longrightarrow C_{10}H_{22}(g) + C_2H_4(g)$$

Another possibility is:

$$C_{12}H_{26}(g) \rightarrow C_8H_{18}(g) + C_4H_8(g)$$

There is a further cracking process which is more versatile, called **thermal cracking**. Thermal cracking is carried out at a higher temperature than catalytic cracking, 800–850 °C. This process is more expensive owing to the higher temperature used. However, larger alkane molecules can be more successfully cracked using this process than by the catalytic method.

Note that in these reactions hydrogen may also be formed during cracking. The amount of hydrogen produced depends on the conditions used. Since smaller hydrocarbons are generally in greater demand than the larger ones, cracking is used to match demand (Table 14.3).

Table 14.3 Percentages of the fractions in crude oil and the demand for them.

Fraction	Approx % in crude oil	Approx % demand
Refinery gas	2	5
Gasoline	21	28
Kerosene	13	8
Diesel oil	17	25
Fuel oil and bitumen	47	34

This means that oil companies are not left with large surpluses of fractions containing the larger molecules.

Figure 14.11 shows the simple apparatus that can be used to carry out cracking reactions in the laboratory. You will notice that in the laboratory we may use a catalyst of broken, unglazed pottery.

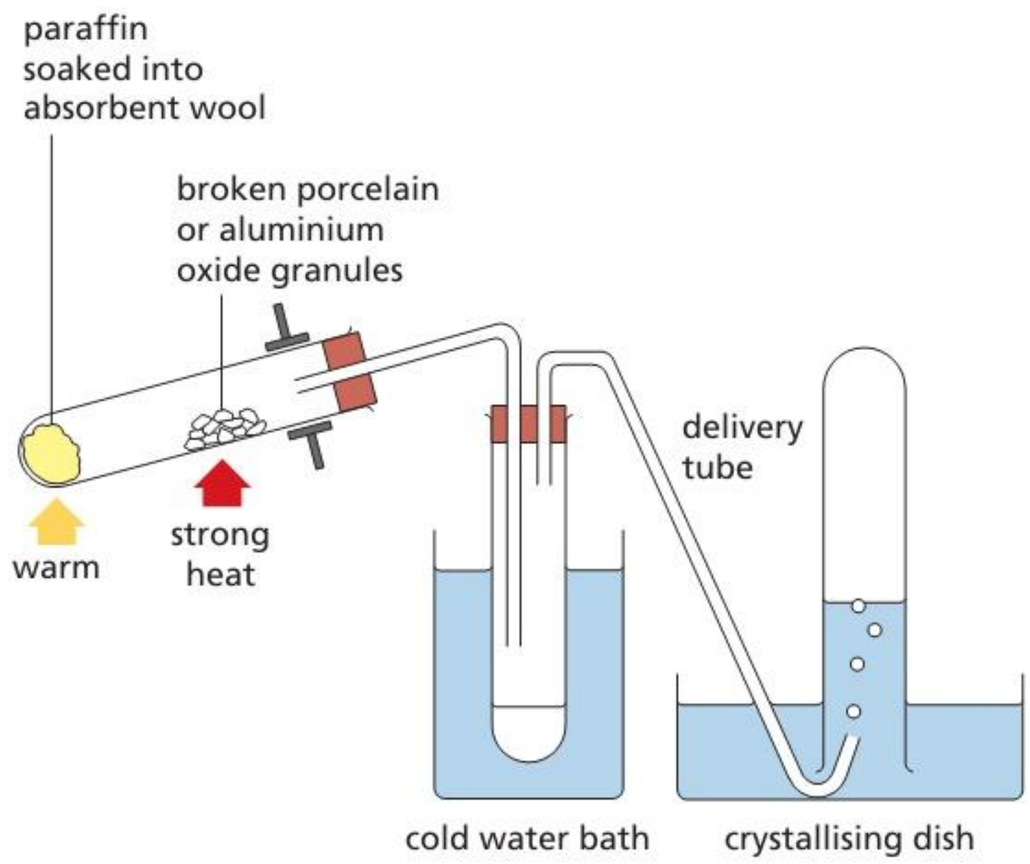


Figure 14.11 The cracking of an alkane in the laboratory.

Questions

1 Using the information in Table 14.2 (p. 223), make an estimate of the boiling point of pentene.
2 Write a balanced chemical equation to represent the process that takes place when decane is cracked.

The chemical behaviour of alkenes

The double bond makes alkenes more reactive than alkanes in chemical reactions. For example, hydrogen adds across the double bond of ethene, under suitable conditions, forming ethane (Figure 14.12).

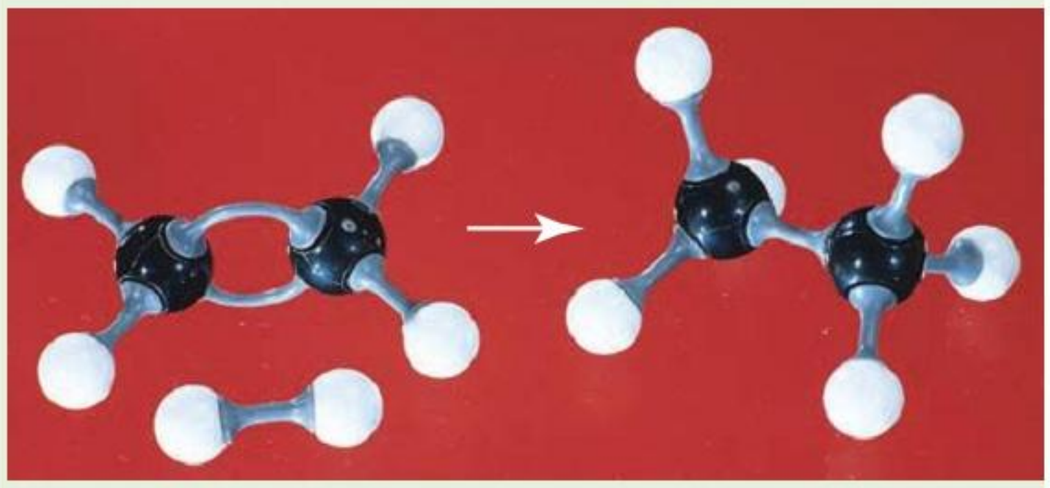


Figure 14.12 The addition of hydrogen to ethene using molecular models.

Addition reactions

Hydrogenation

This reaction is called **hydrogenation**. The conditions necessary for this reaction to take place are a temperature of 200 °C in the presence of a nickel or platinum catalyst.

$$\text{ethene} + \text{hydrogen} \longrightarrow \text{ethane}$$
$$C_2H_4(g) + H_2(g) \longrightarrow C_2H_6(g)$$

Hydrogenation reactions like the one shown with ethene are used in the manufacture of margarines from vegetable oils. Vegetable oils contain fatty acids, such as linoleic acid ($C_{18}H_{32}O_2$). These are unsaturated molecules, containing several double bonds. These double bonds make the molecule less flexible. Hydrogenation can convert these molecules into more saturated ones. Now the molecules are less rigid and can flex and twist more easily, and hence pack more closely together. This in turn causes an increase in the intermolecular forces and so raises the melting point. The now solid margarines can be spread on bread more easily than liquid oils.

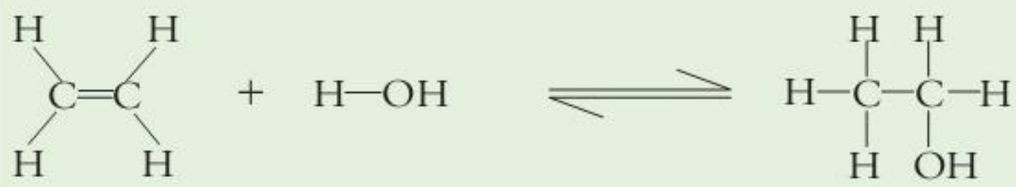
There is another side to this process. Many doctors now believe that unsaturated fats are healthier than saturated ones. Because of this, many margarines are left partially unsaturated. They do not have all the C═C taken out of the fat molecules. However, the matter is far from settled and the debate continues.

Hydration

Another important **addition reaction** is the one used in the manufacture of ethanol. Ethanol has important uses as a solvent and a fuel (p. 94). It is formed when water (as steam) is added across the double bond in ethene. For this reaction to take place, the reactants have to be passed over a catalyst of phosphoric(v) acid (absorbed on silica pellets) at a temperature of 300 °C and pressure of 60 atmospheres (1 atmosphere = 1×10^5 pascals).

$$\text{ethene} + \text{steam} \underset{\text{phosphoric(v) acid catalyst}}{\overset{300\,^\circ\text{C, 60 atm}}{\rightleftharpoons}} \text{ethanol}$$

$$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(g)$$



This reaction is reversible as is shown by the equilibrium (⇌) sign. The conditions have been chosen to ensure the highest possible yield of ethanol. In other words, the conditions have been chosen so that they favour the forward reaction.

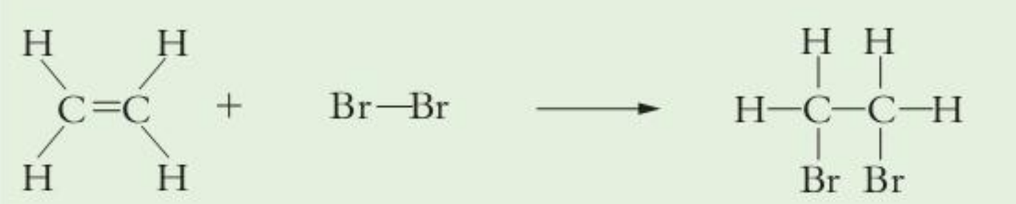
For a further discussion of ethanol and alcohols generally see p. 233.

Halogenation – a test for unsaturated compounds

The addition reaction between bromine dissolved in an organic solvent, or water, and alkenes is used as a chemical test for the presence of a double bond between two carbon atoms. When a few drops of this bromine solution are shaken with the hydrocarbon, if it is an alkene, such as ethene, a reaction takes place in which bromine joins to the alkene double bond. This results in the bromine solution losing its red/brown colour. If an alkane, such as hexane, is shaken with a bromine solution of this type, no colour change takes place (Figure 14.13). This is because there are no double bonds between the carbon atoms of alkanes.

$$\text{ethene} + \text{bromine} \longrightarrow \text{dibromoethane}$$

$$C_2H_4(g) + Br_2\,(\text{in solution}) \longrightarrow C_2H_4Br_2(\text{in solution})$$



Questions

1 What is meant by the term 'addition reaction'?
2 Write a word and balanced chemical equation for the reaction between ethene and hydrogen chloride.
3 Write the structural formula for pentene.

4 Which of the following organic chemicals are alkanes or alkenes?
Propene, C_3H_6
Butanol, C_4H_9OH
Octane, C_8H_{18}
Nonane, C_9H_{20}
Methanoic acid, HCOOH
Butene, C_4H_8
State why you have chosen your answers.

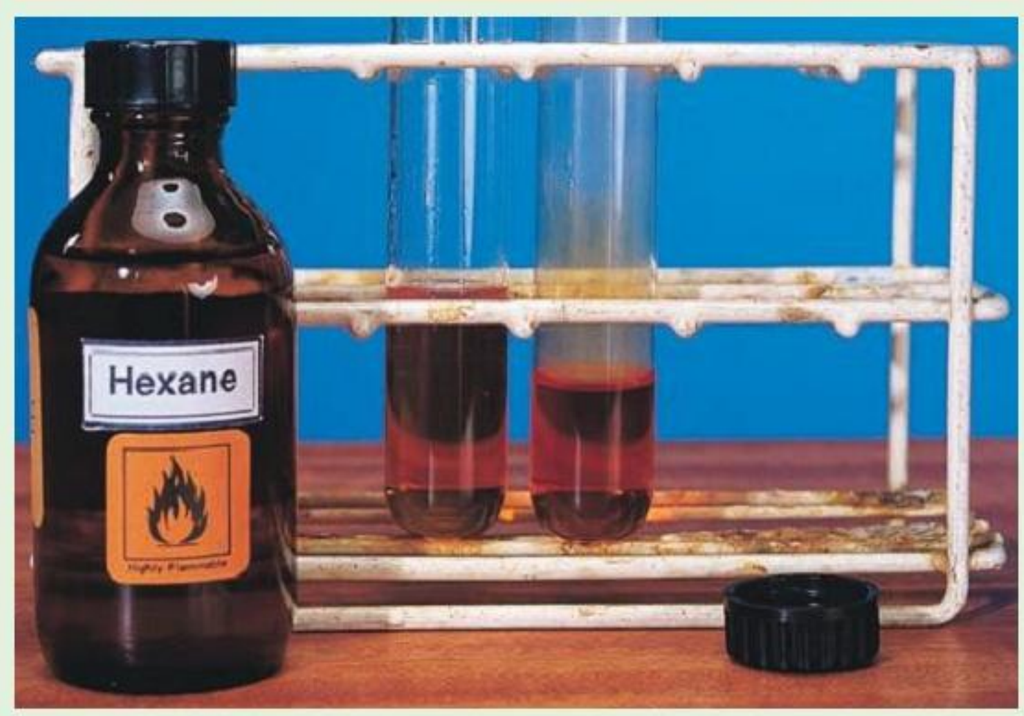
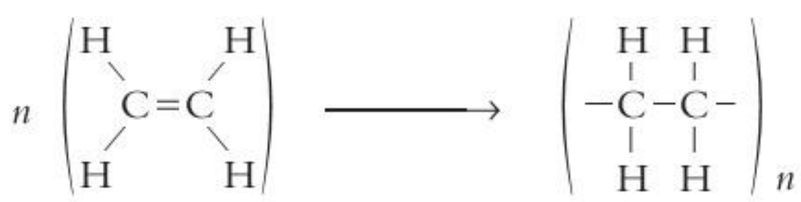


Figure 14.13 The alkene decolourises bromine in 1,1,1-trichloroethane.

A special addition reaction of alkene molecules

Polythene is a plastic that was discovered by accident. Through the careful examination of this substance, when it was accidentally discovered, the plastics industry was born. Polythene is now produced in millions of tonnes worldwide every year. It is made by heating ethene to a relatively high temperature under a high pressure in the presence of a catalyst.



where n is a very large number. In poly(ethene) the ethene molecules have joined together to form a very long hydrocarbon chain (Figure 14.14). The ethene molecules are able to form chains like this because they possess carbon–carbon double bonds.

Figure 14.14 This model shows part of the poly(ethene) polymer chain.

Poly(ethene) is produced in three main forms:

- low density poly(ethene) (LDPE)
- linear low density poly(ethene) (LLDPE)
- high density poly(ethene) (HDPE).

The world production of all types of poly(ethene) is in excess of 52 million tonnes per year.

Poly(ethene) has many useful properties:

- it is easily moulded
- it is an excellent electrical insulator
- it does not corrode
- it is tough
- it is not affected by the weather
- it is durable.

It can be found as a substitute for natural materials in plastic bags, sandwich boxes, washing-up bowls, wrapping film, milk-bottle crates and washing-up liquid bottles (Figure 14.15).

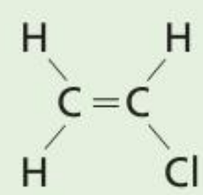
Figure 14.15 These crates are made from poly(ethene).

Other alkene molecules can also produce substances like poly(ethene); for example, propene produces poly(propene), which is used to make ropes and packaging.

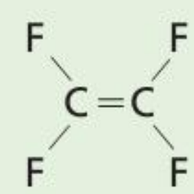
When small molecules like ethene join together to form long chains of atoms, called **polymers**, the process is called **polymerisation**. The small molecules, like ethene, which join together in this way are called **monomers**. A polymer chain, a very large molecule or a macromolecule, often consists of many thousands of monomer units and in any piece of plastic there will be many millions of polymer chains. Since in this polymerisation process the monomer units add together to form only one product, the polymer, the process is called **addition polymerisation**.

Other addition polymers

Many other addition polymers have been produced. Often the plastics are produced with particular properties in mind, for example PVC (polyvinyl chloride or poly(chloroethene)) and PTFE (poly(tetrafluoroethene)). Both of these plastics have monomer units similar to ethene.

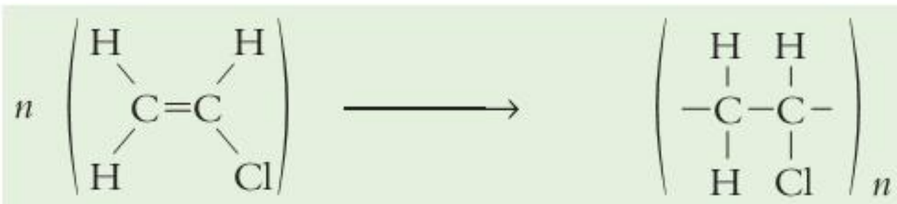


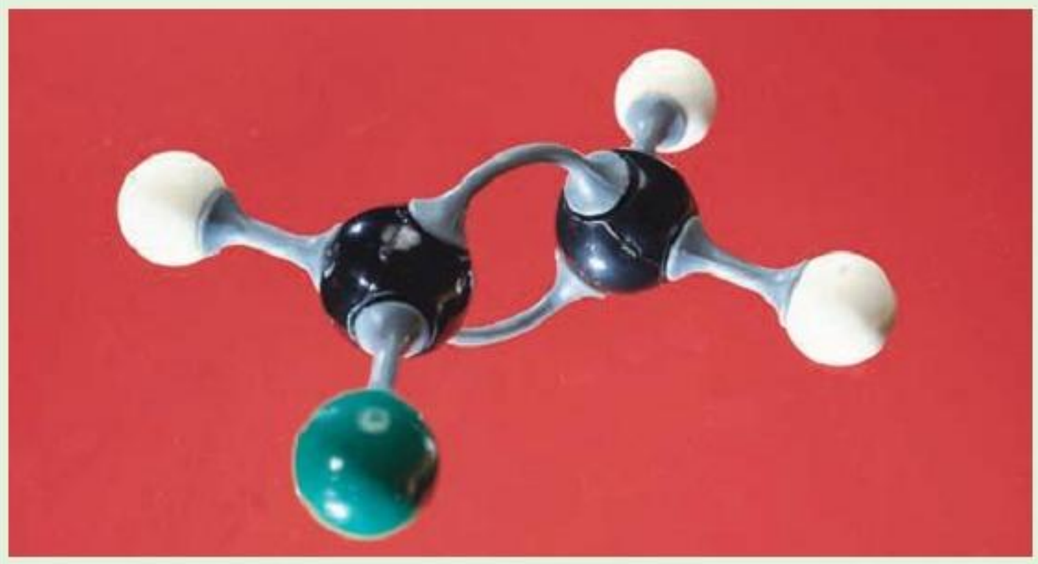
PVC monomer (vinyl chloride or chloroethene)



PTFE monomer (tetrafluoroethene)

If we use chloroethene (Figure 14.16a), the polymer we make is slightly stronger and harder than poly(ethene) and is particularly good for making pipes for plumbing (Figure 14.17).





a Model of chloroethene, the PVC monomer.

b Model of part of a PVC polymer chain.

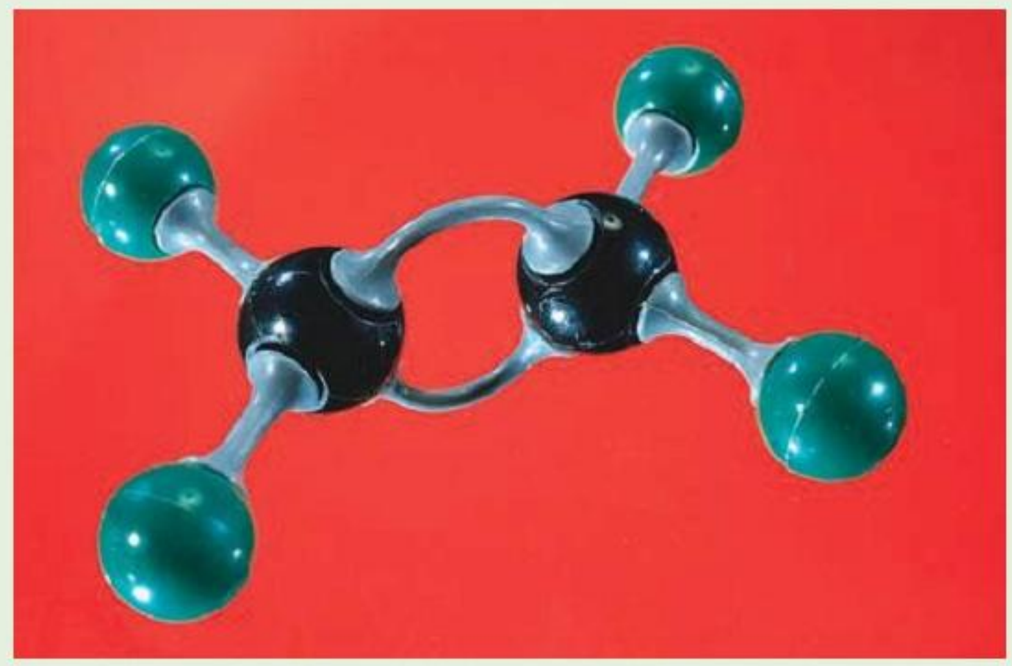
Figure 14.16

Figure 14.17 These pipes are made from PVC.

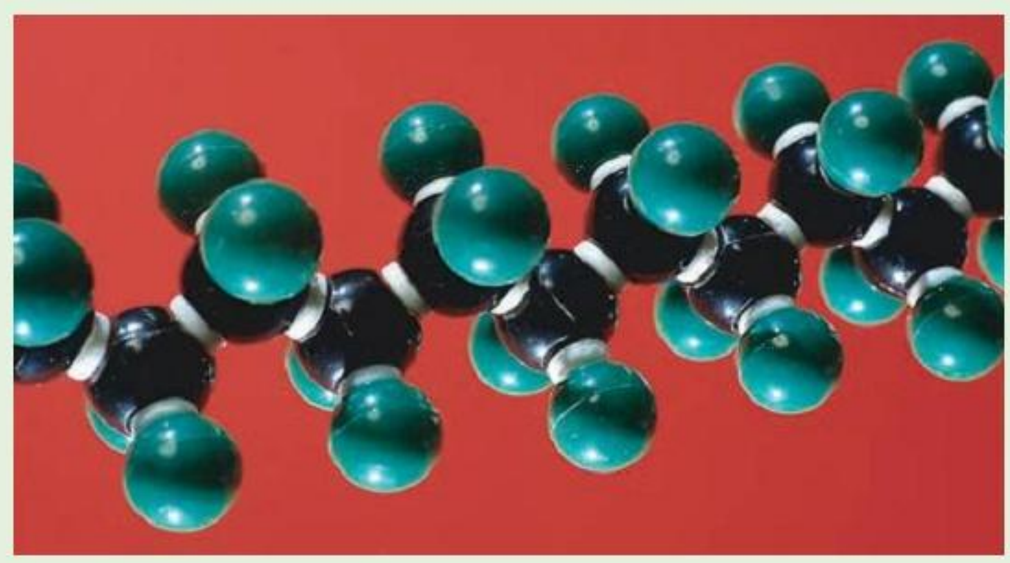
PVC is the most versatile plastic and is the second most widely used, after poly(ethene). Worldwide more than 27 million tonnes are produced annually.

If we start from tetrafluoroethene (Figure 14.18a) the polymer we make, PTFE, has some slightly unusual properties:

- it will withstand very high temperatures, of up to 260 °C
- it forms a very slippery surface
- it is hydrophobic (water repellent)
- it is highly resistant to chemical attack.



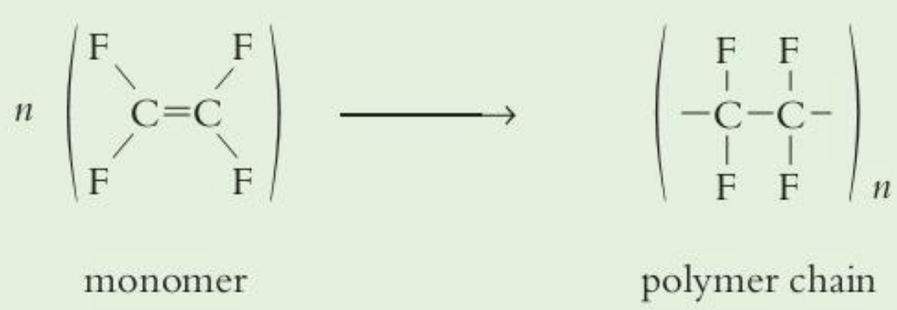
a Model of tetrafluoroethene, the PTFE monomer.



b Model of part of the PTFE polymer chain.

Figure 14.18

These properties make PTFE an ideal 'non-stick' coating for frying pans and saucepans. Every year more than 50 000 tonnes of PTFE are made.



The properties of some addition polymers along with their uses are given in Table 14.4.

Table 14.4 Some addition polymers.

Plastic	Monomer	Properties	Uses
Poly(ethene)	$CH_2{=}CH_2$	Tough, durable	Carrier bags, bowls buckets, packaging
Poly(propene)	$CH_3CH{=}CH_2$	Tough, durable	Ropes, packaging
PVC	$CH_2{=}CHCl$	Strong, hard (less flexible than poly(ethene))	Pipes, electrical insulation, guttering
PTFE	$CF_2{=}CF_2$	Non-stick surface, withstands high temperatures	Non-stick frying pans, soles of irons
Polystyrene	$CH_2{=}CHC_6H_5$	Light, poor conductor of heat	Insulation, packaging (especially as foam)
Perspex	$CH_2{=}C(CO_2CH_3)CH_3$	Transparent	Used as a glass substitute

Thermosoftening and thermosetting plastics

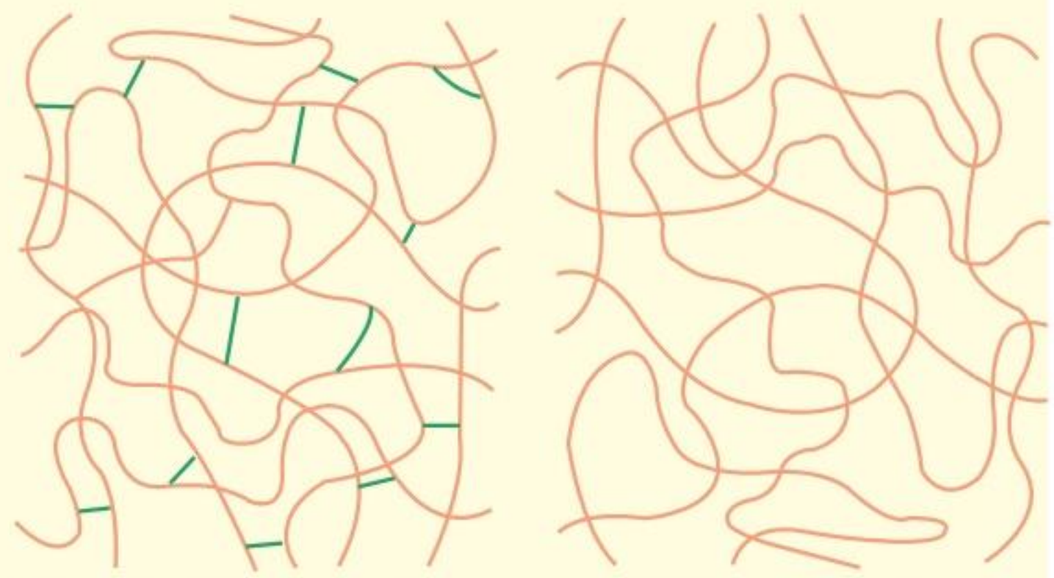
Plastics can be put into one of two categories. If they melt or soften when heated (like poly(ethene), PVC and polystyrene) then they are called **thermoplastics** or **thermosoftening plastics**. If they do not soften on heating but only char and decompose on further heating, they are known as **thermosetting plastics**.

Thermoplastics are easily moulded or **formed** into useful articles. Once they are molten they can be injected or blown into moulds, and a variety of different-shaped items can be produced (Figure 14.19). Thermosetting plastics can be heated and moulded only once, usually by compression moulding (Figure 14.20).

Figure 14.19 These objects are made from thermosoftening plastics.

Figure 14.20 These objects are made from compression-moulded thermosetting plastics.

Figure 14.21 shows the different molecular structures for thermosetting and thermosoftening plastics. Thermosetting plastics have polymer chains which are linked or bonded to each other to give a **cross-linked** structure, and so the chains are held firmly in place and no softening takes place on heating. Thermosoftening plastics do not have polymer chains joined in this way, so when they are subjected to heat their polymer chains flow over one another and the plastic softens.



a In thermosetting plastic the chains are cross-linked.

b In thermosoftening plastic there is no cross-linking.

Figure 14.21

Disposal of plastics

In the last 30 to 40 years plastics have taken over as replacement materials for metals, glass, paper and wood as well as for natural fibres such as cotton and wool. This is not surprising since plastics are light, cheap, relatively unreactive, can be easily moulded and can be dyed bright colours. However, plastics have contributed significantly to the household waste problem, up to 10% in some countries, and it's getting worse (Figure 14.22)!

Figure 14.22 This plastic waste is ready to go to landfill.

In the recent past, much of our plastic waste has been used to landfill disused quarries. However, all over the world these sites are getting harder to find and it is becoming more and more expensive. The alternatives to dumping plastic waste are certainly more economical and more satisfactory.

- Incineration schemes have been developed to use the heat generated for heating purposes (Figure 14.23). However, problems with the combustion process (which can result in the production of toxic gases) mean that especially high temperatures have to be employed during the incineration process.
- Recycling – large quantities of black plastic bags and sheeting are produced for resale.
- Biodegradable plastics, as well as those polymers that degrade in sunlight (**photodegradable**, Figure 14.24a), have been developed. Other common categories of degradable plastics include synthetic biodegradable plastics which are broken down by bacteria, as well as plastics which dissolve in water (Figure 14.24b). The property that allows plastic to dissolve in water has been used in relatively new products, including soluble capsules containing liquid detergent. However, the vast majority of polymers are still non-biodegradable.

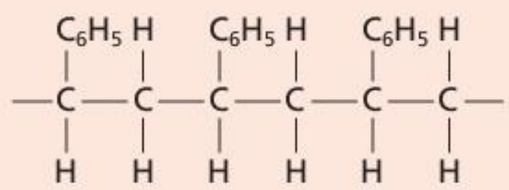
Figure 14.23 A plastic incineration plant in Germany.

a This plastic bag is photodegradable.

b This plastic dissolves in water.
Figure 14.24

Questions

1 Write the general equation to represent the formation of polystyrene from its monomer.
2 Give two advantages and two disadvantages of plastic waste (rubbish).
3 Draw the structure of the repeating unit of the addition polymer formed from $CH_3—CH{=}CH_2$.
4 Draw the structure of the monomer from which the addition polymer below has been produced.



Checklist

After studying Chapter 14 you should know and understand the following terms.

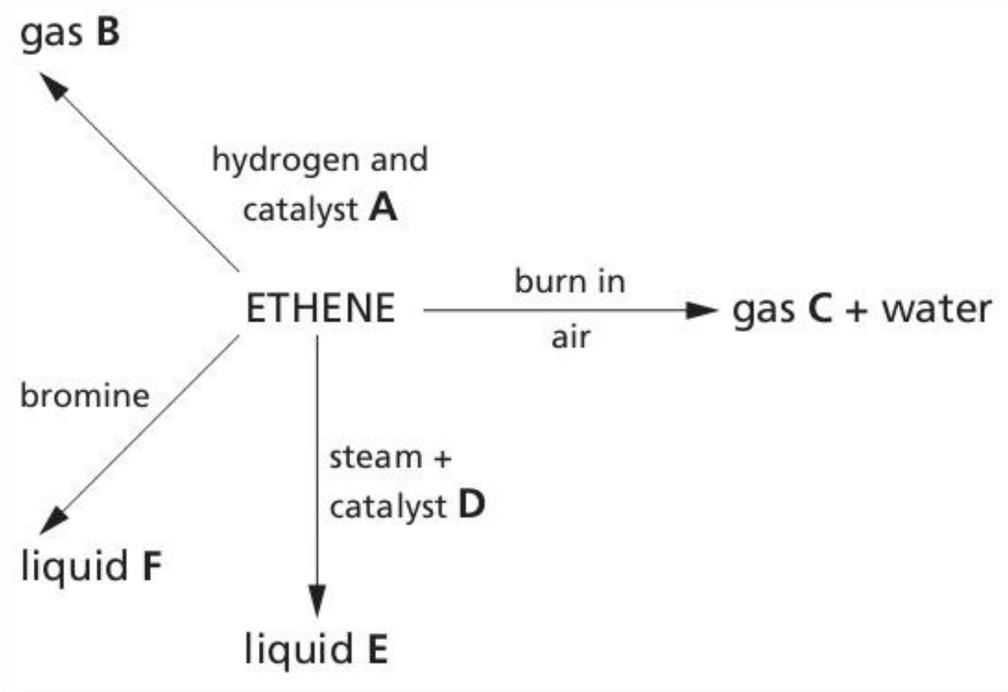
- **Addition polymer** A polymer formed by an addition reaction. For example, poly(ethene) is formed from ethene.
- **Addition reaction** A reaction in which an atom or group of atoms is added across a carbon-carbon double bond.
- **Alkanes** A family of saturated hydrocarbons with the general formula C_nH_{2n+2}. The term 'saturated', in this context, is used to describe molecules that have only single bonds. The alkanes can only undergo substitution reactions in which there is replacement of one atom in the molecule by another atom.
- **Alkenes** A family of unsaturated hydrocarbons with the general formula C_nH_{2n}. The term 'unsaturated', in this context, is used to describe molecules which contain one or more carbon–carbon double bonds. Unsaturated compounds undergo addition reactions across the carbon–carbon double bonds and so produce saturated compounds. The addition of hydrogen across the carbon–carbon double bonds is used to reduce the amount of unsaturation during the production of margarines.
- **Biodegradable plastics** Plastics designed to degrade (decompose) under the influence of bacteria.
- **Catalytic cracking** The decomposition of higher alkanes into alkenes and alkanes of lower relative molecular mass. The process involves passing the larger alkane molecules over a catalyst of aluminium and chromium oxides, heated to 500°C.
- **CFC** Abbreviation for chlorofluorocarbon, a type of organic compound in which some or all of the hydrogen atoms of an alkane have been replaced by fluorine and chlorine atoms. These substances are generally unreactive but they can diffuse into the stratosphere where they break down under the influence of ultraviolet light. The products of this photochemical process then react with ozone (in the ozone layer). Because of this, their use has been discouraged. They are now being replaced by hydrochlorofluorocarbons (HCFCs).
- **Chain reaction** A reaction which is self-sustaining owing to the products of one step of the reaction assisting in promoting further reaction.
- **Cross-linking** The formation of side covalent bonds linking different polymer chains and therefore increasing the rigidity of, say, a plastic. Thermosetting plastics are usually heavily cross-linked.
- **Free radicals** Atoms or groups of atoms with unpaired electrons and which are therefore highly reactive. They can be produced by high-energy radiation such as ultraviolet light in photochemical reactions.
- **Halogenoalkanes** Organic compounds in which one or more hydrogen atoms of an alkane have been substituted by halogen atoms such as chlorine.
- **Homologous series** A series of compounds in which each member differs from the next by a specific number and kind of atom. These compounds have the same general formula and similar properties.
- **Isomers** Compounds which have the same molecular formula but different structural arrangements of the atoms.
- **Monomer** A simple molecule, such as ethene, which can be polymerised.
- **Organic chemistry** The branch of chemistry concerned with compounds of carbon found in living organisms.
- **Photodegradable plastics** Plastics designed to degrade under the influence of sunlight.
- **Polymer** A substance possessing very large molecules consisting of repeated units or monomers. Polymers therefore have a very large relative molecular mass.
- **Polymerisation** The chemical reaction in which molecules (monomers) join together to form a polymer.
- **Saturated hydrocarbon** A type of hydrocarbon molecule in which the molecule has the maximum possible number of hydrogen atoms and so has no double bonds.
- **Substitution reaction** A reaction in which an atom or group of atoms is replaced by another atom or group of atoms.
- **Test for unsaturation** A few drops of bromine dissolved in an organic solvent are shaken with the hydrocarbon. If it is decolourised, the hydrocarbon is unsaturated.
- **Thermal cracking** The decomposition of higher alkanes to alkenes of lower relative molecular mass at high temperatures, 800–850°C.
- **Thermoplastics** Plastics which soften when heated (for example poly(ethene), PVC).
- **Thermosetting plastics** Plastics which do not soften on heating but only char and decompose (for example, Bakelite and melamine).
- **Unsaturated hydrocarbon** A hydrocarbon molecule which contains double or triple covalent bonds between carbon atoms.

Organic chemistry 1

Additional questions

1 Explain the following.
   a Ethene is called an unsaturated hydrocarbon.
   b The cracking of larger alkanes into simple alkanes and alkenes is important to the petrochemical industry.
   c The conversion of ethene to ethanol is an example of an addition reaction.

2 The following question is about some of the reactions of ethene.



   a Give the names and formulae for substances **A** to **F**.
   b (i) Write a word and balanced chemical equation to represent the reaction in which liquid **E** is formed.
      (ii) What reaction conditions are required for the process to take place?
      (iii) Hydrogen is used in the production of margarine to remove unsaturation. Explain what you understand by this statement.
   c Name the homologous series that gas **B** belongs to.
   d Describe a chemical test which would allow you to identify gas **C**.

3 a Crude oil is a mixture of *hydrocarbons* which belong to the *homologous series* called the *alkanes*. This mixture can be separated into fractions by the process of *fractional distillation*. Some of the fractions obtained are used as *fuels*. Some of the other fractions are subjected to *catalytic cracking* in order to make *alkenes*.
   Explain the meaning of the terms in italics.
   b Alkanes can be converted into substances which are used as solvents. To do this the alkane is reacted with a halogen, such as chlorine, in the presence of ultraviolet light.
      (i) Write a word and balanced chemical equation for the reaction between methane and chlorine.
      (ii) Name the type of reaction taking place.
      (iii) Highly reactive chlorine atoms are produced in the presence of ultraviolet light. When atoms are produced in this way, what are they called?
      (iv) Write a balanced chemical equation for the reaction which takes place between CHF_3 and Cl_2 to produce a chlorofluorocarbon (CFC).
      (v) Why are CFCs such a problem?

4 a Ethene, C_2H_4, is the starting material for making plastic carrier bags.

      (i) Name the type of chemical change taking place in the diagram above.
      (ii) Name the product formed by this reaction.
      (iii) The alkene, ethene, is made by cracking large alkane molecules. Describe a simple chemical test to show that ethene is present.
   b The majority of carrier bags are difficult to dispose of.
      (i) Explain why carrier bags should not just be thrown away.
      (ii) Explain why the majority of plastic carrier bags are recycled.
      (iii) Give one advantage that a plastic carrier bag has over one made out of paper.
   c A label like the one below is found on some plastic carrier bags.

> This plastic carrier bag is made from a substance that is made from the chemical elements carbon and hydrogen only. When the carrier bag is burned it produces carbon dioxide and water. These substances are natural and will not harm the environment.

(i) What is the meaning of the term *element*?
(ii) What is the name given to the type of compound that contains the elements carbon and hydrogen only?
(iii) When the plastic bag burns, heat energy is given out. What name is used to describe reactions that give out heat energy?
(iv) The plastic bag will probably give off a toxic gas when it is burned. Why is this the case?

5 Alkanes and alkenes are hydrocarbons. They are composed of molecules which contain covalent bonds. For each of the molecules below, use a dot and cross diagram to show the bonding it contains.
a Methane, CH_4.
b Propene, C_3H_6.
c Propane, C_3H_8.
d Ethene, C_2H_4.

6 Crude oil is an important source of organic chemical fuels. It is refined by fractional distillation. Use the information in the table below to answer the questions which follow.

Fraction	Boiling point/°C
A	40
B	80
C	200
D	350
E	above 350

a For each of the questions that follow, give the letter of the fraction which is most appropriate as an answer. You should also give a reason for your answer in each case.
(i) Which fraction would contain the most volatile substances?
(ii) Which of the fractions would collect at the bottom of the fractionating column?
(iii) Which fraction could be used as a fuel for cars?
(iv) Which fraction would contain the largest molecules?
b Some of the fractions undergo a further process called cracking to produce further substances.
(i) Explain what you understand by the term 'cracking'. What conditions are employed when cracking occurs?
(ii) Write a word and balanced chemical equation to show how octane can be produced by the cracking of $C_{15}H_{32}$.

7 a A hydrocarbon contains 92.3% by mass of carbon. Work out the empirical formula of this hydrocarbon.
b The relative molecular mass of this hydrocarbon was found by mass spectrometry to be 78. Work out its molecular formula. (A_r: H = 1, C = 12)

8 a Which of the following formulae represent alkanes, which represent alkenes and which represent neither?

CH_3, C_6H_{12}, C_5H_{12}, C_6H_6, C_9H_{20}, $C_{12}H_{24}$, $C_{20}H_{42}$, C_2H_4, C_8H_{18}, C_3H_7

b Draw all the possible isomers which have the molecular formula C_6H_{14}.

15 Organic chemistry 2

Alcohols (R—OH)
Combustion
Oxidation
Cholesterol – a complex molecule that contains the –OH group

Biotechnology
Baking – an important use of biotechnology
New applications of biotechnology

Carboxylic acids
Other carboxylic acids

Soaps and detergents
Soapless detergents
Manufacture of soapless detergents

Condensation polymers

Some biopolymers
Starch
Amino acids
DNA

Pharmaceuticals
Drug abuse

Checklist

Additional questions

In Chapter 14 we discussed the organic compounds that are obtained from oil, the alkanes, and how they are converted into another homologous series of hydrocarbons called the alkenes. If you replace one of the hydrogen atoms on an alkane molecule with a group such as –OH, the hydroxyl group, then you get a new homologous series called the alcohols. If you replace one of the hydrogen atoms on an alkane molecule with a –COOH group (Figure 15.1) then you get a homologous series called the carboxylic acids or alkanoic acids. Whichever group you have attached, it will bring with it a new set of physical and chemical properties. These groups are known as **functional groups**. The functional group is the group of atoms responsible for the characteristic reactions of the organic compound.

Table 15.1 shows some examples of functional groups. In the table, R represents an alkyl group or a hydrogen atom. An alkyl group has the general formula based on the alkanes, i.e. C_nH_{2n+1}. When $n = 1$, R = CH_3; when $n = 2$, R = C_2H_5, and so on.

Figure 15.1 This fruit juice contains plenty of vitamin C or ascorbic acid, which contains the functional group –COOH.

Table 15.1 The functional groups present in some homologous series of organic compounds.

Class of compound	Functional group
Alcohols	R—OH
Carboxylic acids	R—COOH
Esters	R—COOR
Halogenoalkanes	R—X (where X is a halogen atom such as Cl, Br, etc.)

Alcohols (R—OH)

The alcohols (alkanols) form another homologous series with the general formula $C_nH_{2n+1}OH$ (or R—OH, where R represents an alkyl group). All the alcohols possess an –OH as the functional group. Table 15.2 shows the names and condensed formulae of the first four members along with their melting and boiling points.

Table 15.2 Some members of the alcohol family.

Alcohol	Formula	Melting point/ °C	Boiling point/ °C
Methanol	CH_3OH	–94	64
Ethanol	CH_3CH_2OH	–117	78
Propanol	$CH_3CH_2CH_2OH$	–126	97
Butanol	$CH_3CH_2CH_2CH_2OH$	–89	117

Figure 15.2 shows the arrangement of the atoms in the fully displayed (structural) formulae of these members of this family or homologous series.

The alcohols are named by reference to the corresponding alkane (see Chapter 14, p. 219), always ensuring that the hydrocarbon chain is numbered from the end that gives the lowest

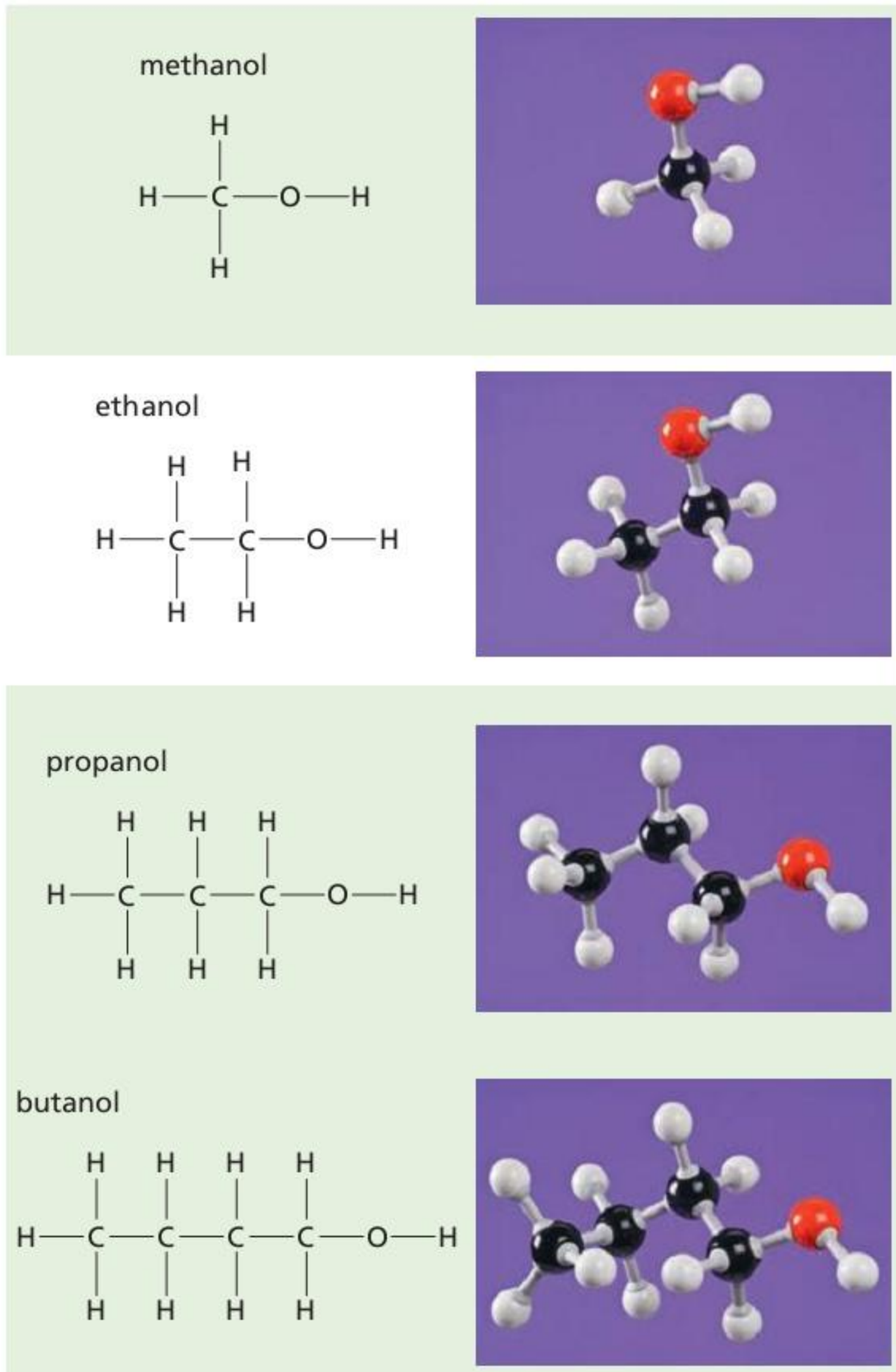


Figure 15.2 The molecules look like their corresponding models in the photographs.

number to the position of the –OH group. If the –OH group is positioned at the end of the alcohol, for example in $CH_3CH_2CH_2CH_2OH$, then the position of this group is shown by numbering the carbon atom it is attached to as '1'. So this molecule is called butan-1-ol, or butanol for short. This type of alcohol is called a primary alcohol.

Alcohols have high boiling points and relatively low volatility. Alcohol molecules are like water molecules (H—OH) in that they are polar (see Chapter 11, p. 186).

Alcohol molecules are polar because of the presence of the –OH group, in which the hydrogen attached to oxygen creates (within this group) a small difference in charge (Figure 15.3). Other organic molecules that are polar are carboxylic acids such as ethanoic acid (p. 237).

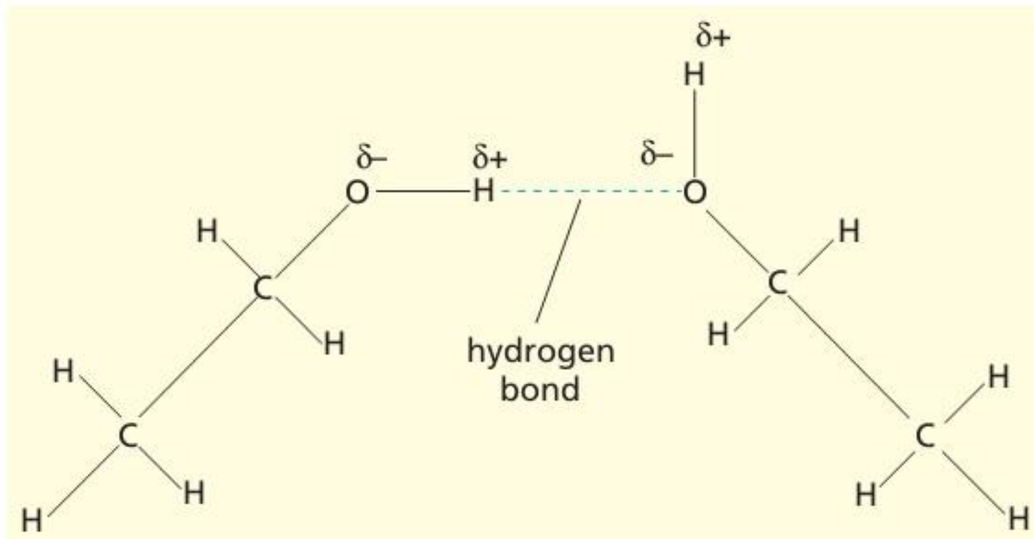


Figure 15.3 Polar –OH groups in ethanol molecules lead to hydrogen bonding and an attraction between neighbouring molecules.

Because of the presence of the polar –OH groups, there is a relatively strong polar attraction between alcohol molecules. This polar attraction between the charges δ+ and δ– in neighbouring molecules is called a hydrogen bond. This means that the molecules have a much stronger attraction for each other than molecules of the corresponding alkane of similar relative molecular mass, M_r. For example, ethanol has an M_r value of 46 and is a liquid at room temperature with a boiling point of 78 °C, whilst propane ($CH_3CH_2CH_3$) has an M_r of 44 and is a gas at room temperature with a boiling point of –42 °C.

It is possible to think of alcohol molecules as water molecules in which an H atom has been replaced by an alkyl group, for example $-C_2H_5$ in C_2H_5OH. This close similarity between their molecules explains why water and alcohols with small molecules, such as methanol and ethanol, are miscible. They mix because of the presence of the polar –OH group in both molecules.

Many other materials, such as food flavourings, are made from ethanol. As ethanol is also a very good solvent and evaporates easily, it is used extensively as a solvent for paints, glues, aftershave and many other everyday products (Figure 15.4).

Figure 15.4 This aftershave contains alcohol.

Ethanol is by far the most important of the alcohols and is often just called 'alcohol'. Ethanol can be produced by fermentation (p. 236) as well as by the hydration of ethene (Chapter 14, p. 224). It is a neutral, colourless, volatile liquid which does not conduct electricity. The more concentrated forms of alcoholic drinks such as the spirits whiskey and brandy contain high concentrations of ethanol. These are produced by distillation after the fermentation is complete (Chapter 2, p. 20).

Combustion

Ethanol burns quite readily with a clean, hot flame.

ethanol + oxygen → carbon dioxide + water + energy

$$CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g) + \text{energy}$$

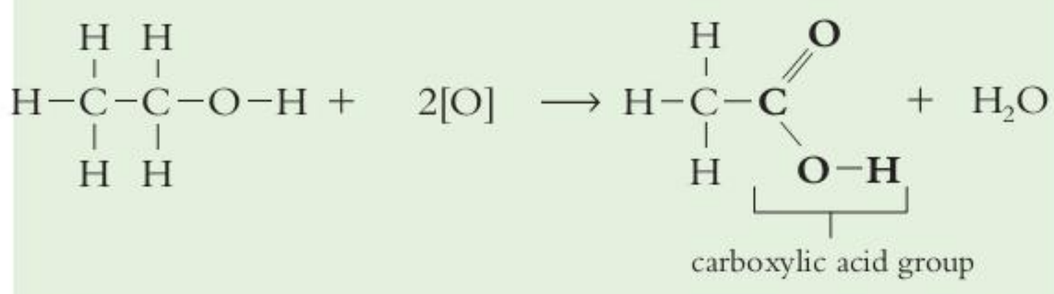
As methylated spirit, it is used in spirit (camping) stoves. Methylated spirit is ethanol with small amounts of poisonous substances added to stop people drinking it. Some countries, like Brazil, already use ethanol mixed with petrol as a fuel for cars (Chapter 6, p. 94) and this use is increasing worldwide.

Oxidation

Ethanol can be oxidised to ethanoic acid (an organic acid also called acetic acid) by powerful oxidising agents, such as warm acidified potassium dichromate(VI), or potassium manganate(VII). During the reaction the orange colour of potassium dichromate(VI) changes to a dark green (Figure 15.5) as the ethanol is oxidised to ethanoic acid.

ethanol + oxygen (from potassium dichromate (VI)) $\xrightarrow{\text{heat}}$ ethanoic acid + water

$$CH_3CH_2OH(l) + 2[O] \longrightarrow CH_3COOH(aq) + H_2O(l)$$



Question

1 Write a word and balanced chemical equation for:
   a the combustion of butanol
   b the oxidation of butanol.

Figure 15.5 Orange potassium dichromate(VI) slowly turns green as it oxidises ethanol to ethanoic acid. Potassium manganate(VII) turns from purple to colourless.

Cholesterol – a complex molecule that contains the –OH group

Cholesterol is a naturally occurring and essential chemical. It belongs to a family of chemicals called steroids and also contains an alcohol group (Figure 15.6). Cholesterol is found in almost all of the tissues in the body, including nerve cells. Levels of cholesterol above normal (above 6.5 mmol/l) are associated with an increased risk of heart disease. Cholesterol hardens and blocks off arteries by building up layers of solid material (atheroma) inside the arteries (Figure 15.7, p. 236). This is particularly serious if the arteries that supply the heart or brain are blocked. Simple tests are now available to monitor cholesterol levels and people with high levels can be treated and can follow special low-fat and low-cholesterol diets.

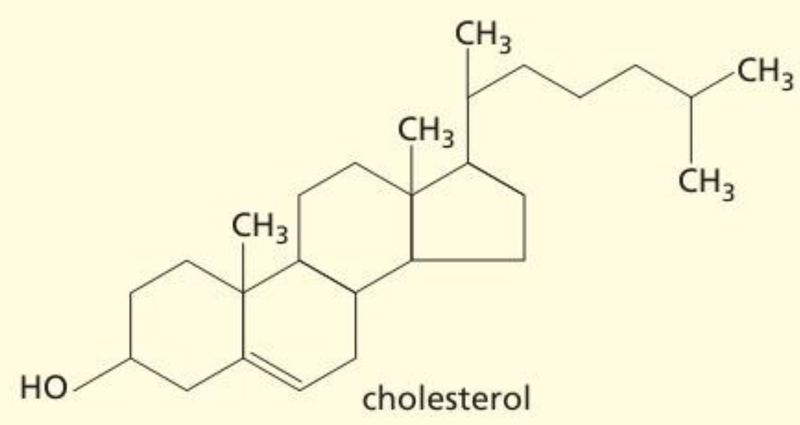
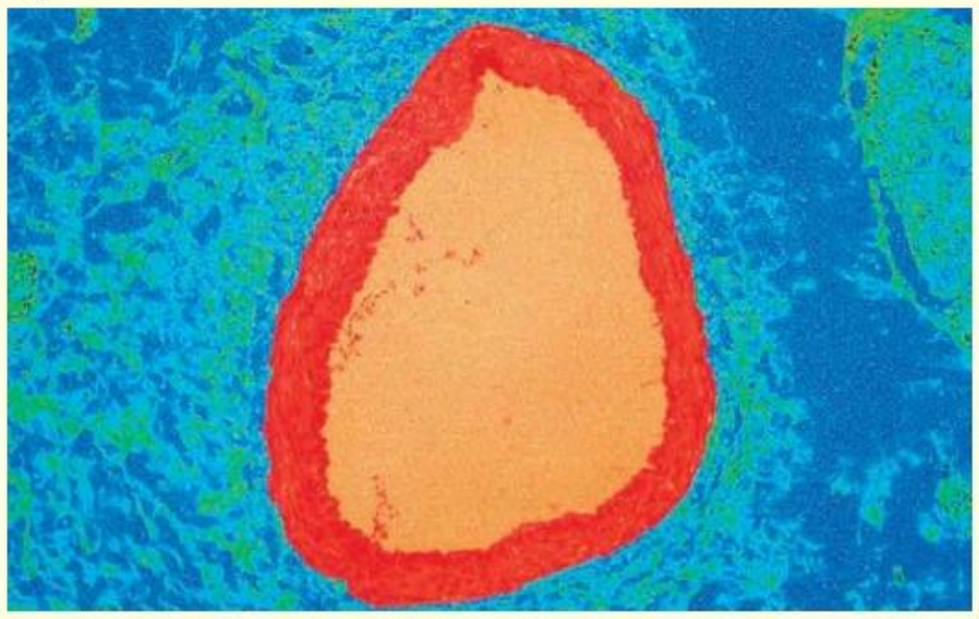


Figure 15.6 The structure of cholesterol.

a This artery is being blocked by atheroma, which may be related to high levels of cholesterol in the blood.



b This is a healthy artery.
Figure 15.7

Biotechnology

Biotechnology involves making use of micro-organisms or their components, such as enzymes, for the benefit of humans to produce, for example, foods such as yoghurt and bread. One of the oldest biotechnologies is that of **fermentation**. It involves a series of biochemical reactions brought about by micro-organisms or enzymes.

Fermentation in the laboratory can be carried out using sugar solution. A micro-organism called yeast is added to the solution. The yeast uses the sugar for energy during **anaerobic respiration** (respiration without oxygen), and so the sugar is broken down to give carbon dioxide and ethanol. The best temperature for this process to be carried out is at 37 °C.

$$\text{glucose} \xrightarrow{\text{yeast}} \text{ethanol} + \text{carbon dioxide}$$

$$C_6H_{12}O_6(aq) \longrightarrow 2C_2H_5OH(l) + 2CO_2(g)$$

Figure 15.8 shows a simple apparatus for obtaining ethanol from glucose in the laboratory.

If a bottle of wine is left exposed to the atmosphere then further oxidation takes place and the ethanol is oxidised to ethanoic acid.

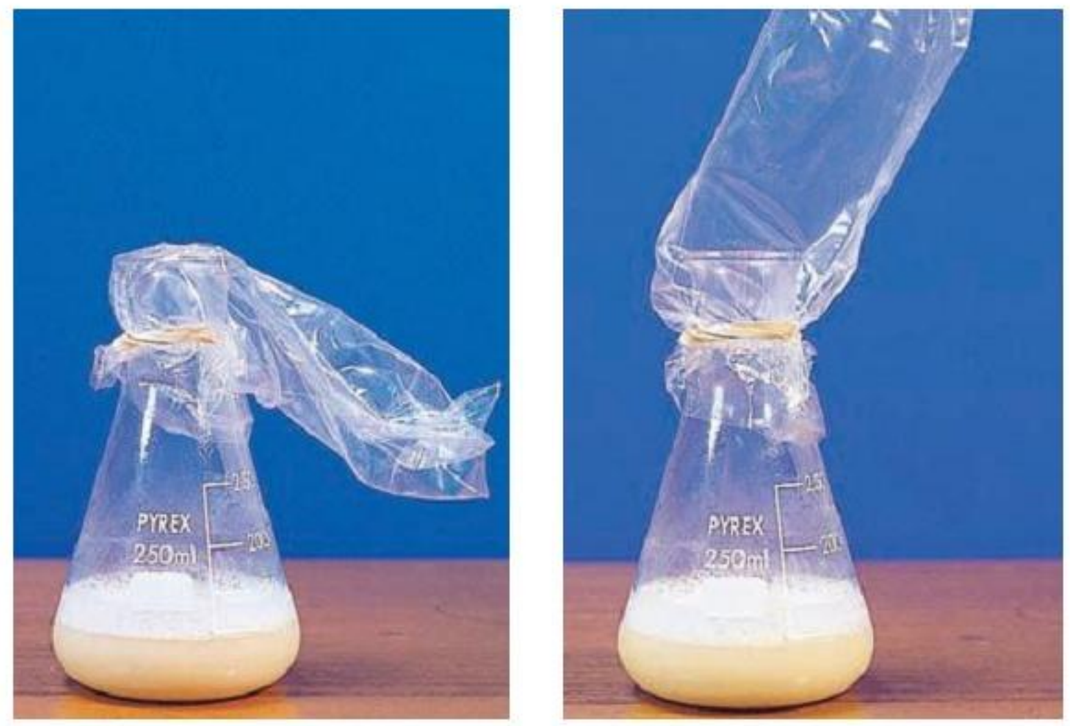


Figure 15.8 Fermenting glucose and yeast to produce ethanol. The bag is inflated during the experiment by CO_2.

Baking – an important use of biotechnology

To make bread, fresh yeast is mixed with warm sugar solution and the mixture added to the flour. This dough mixture is then put into a warm place to rise. The dough rises due to the production of carbon dioxide from **aerobic respiration** (respiration with oxygen) by the yeast. The products of this style of respiration are different to those of anaerobic respiration.

$$\text{sugar} + \text{oxygen} \xrightarrow{\text{yeast}} \text{carbon dioxide} + \text{water} + \text{energy}$$

$$C_6H_{12}O_6(aq) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l) + \text{energy}$$

After the dough has 'risen', it is baked and the heat kills the yeast and the bread stops rising.

New applications of biotechnology

A large number of firms throughout the world are investing large sums of money in the newer biotechnology applications now in use.

- Enzymes can be isolated from micro-organisms and used to catalyse reactions in other processes. For example, proteases are used in biological detergents to digest protein stains such as blood and food. Also, catalase is used in the rubber industry to help convert latex into foam rubber.
- Our ability to manipulate an organism's genes to make it useful to us is called **genetic engineering**. This is being used, for example, to develop novel plants for agriculture as well as making important human proteins such as the hormones insulin and growth hormone.

However, a word of caution is necessary. The new biotechnologies may not be without dangers. For example, new pathogens (organisms that cause disease) might be created accidentally. Also, new pathogens may be created deliberately for use in warfare. As you can imagine, there are very strict guidelines covering these new biotechnologies, especially in the area of research into genetic engineering.

Question

1 What do you understand by the term 'biotechnology'? In your answer make reference to the making of bread.

Carboxylic acids

The carboxylic acids form another homologous series, this time with the general formula $C_nH_{2n+1}COOH$. All the carboxylic acids possess –COOH as their functional group. Table 15.3 shows the first four members of this homologous series along with their melting and boiling points. Figure 15.9 shows the actual arrangement of the atoms in these members of this family.

Table 15.3 Some members of the carboxylic acid series.

Carboxylic acid	Formula	Melting point/ °C	Boiling point/ °C
Methanoic acid	HCOOH	9	101
Ethanoic acid	CH_3COOH	17	118
Propanoic acid	CH_3CH_2COOH	–21	141
Butanoic acid	$CH_3CH_2CH_2COOH$	–6	164

Methanoic acid is present in stinging nettles and ant stings. Ethanoic acid, however, is the most well known as it is the main constituent of vinegar. Like other acids, ethanoic acid affects indicators and will react with metals such as magnesium. However, whereas the mineral acids such as hydrochloric acid are called strong acids, ethanoic acid is a weak acid (Chapter 8, p. 120). Even though it is a weak acid, it will still react with bases to form salts. For example, the salt sodium ethanoate is formed when ethanoic acid reacts with dilute sodium hydroxide.

ethanoic acid + sodium hydroxide → sodium ethanoate + water

$$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$$

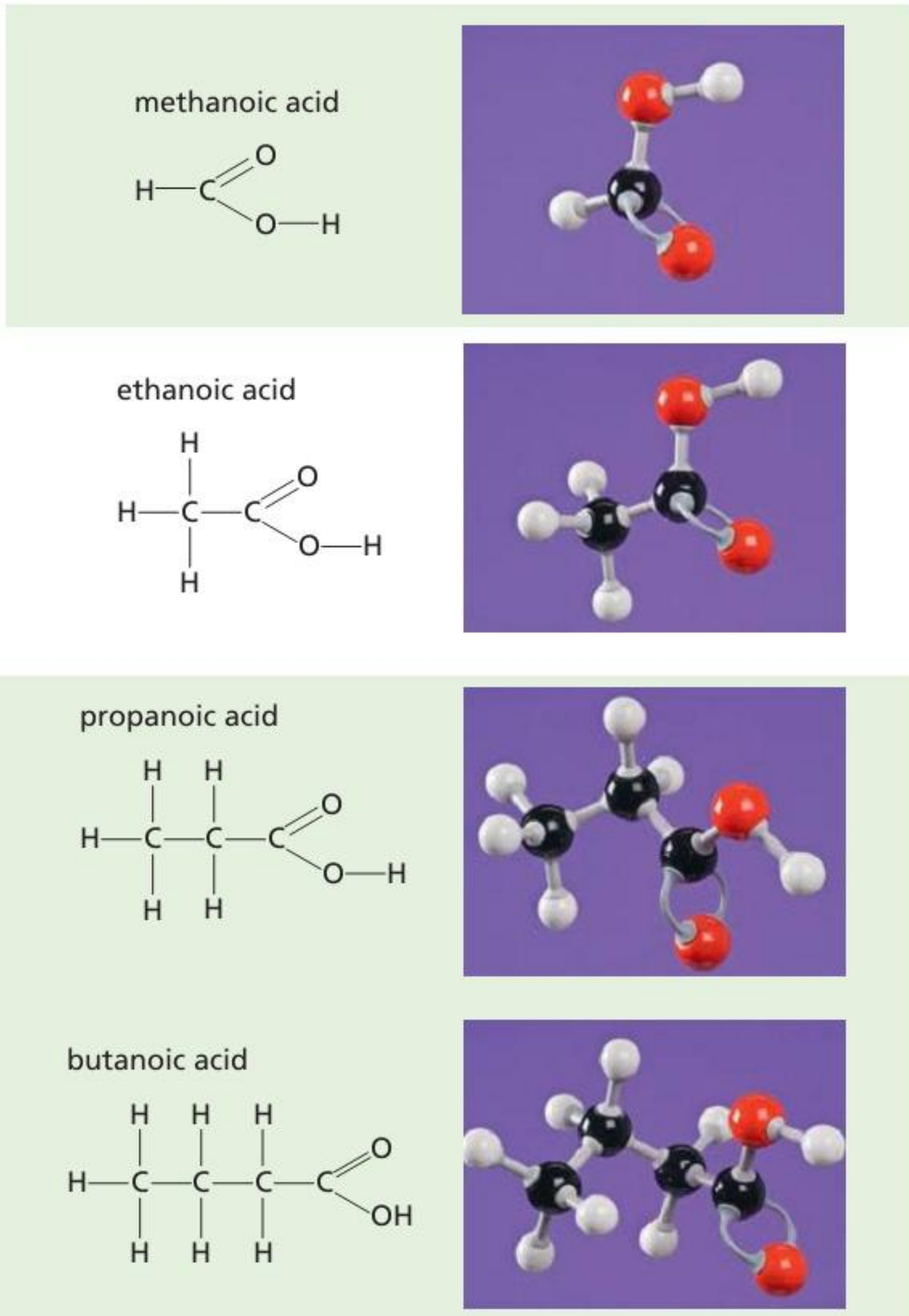


Figure 15.9 The molecules look like the 3D models in the photographs.

Ethanoic acid also undergoes other typical reactions of acids, in that it reacts with indicators, metals and carbonates in the usual way.

Ethanoic acid will react with ethanol, in the presence of a few drops of concentrated sulfuric acid, to produce ethyl ethanoate – an ester.

ethanoic acid + ethanol $\underset{H_2SO_4}{\overset{conc}{\rightleftharpoons}}$ ethyl ethanoate + water

$$CH_3COOH(l) + C_2H_5OH(l) \rightleftharpoons CH_3COOC_2H_5(aq) + H_2O(l)$$

This reaction is called **esterification**.

Members of the 'ester' family have strong and pleasant smells. They have the general formula $C_nH_{2n+1}COOC_xH_{2x+1}$. Esters are named after the acid and alcohol from which they are derived:

- name – alcohol part first, acid part second, e.g. propyl ethanoate
- formula – acid part first, alcohol part second, e.g. $CH_3COOC_3H_7$

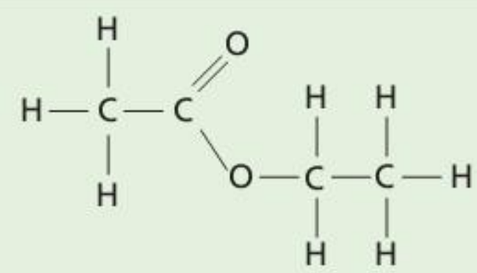
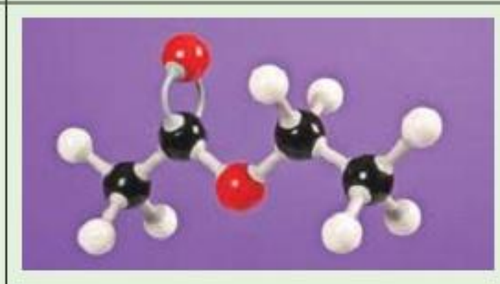
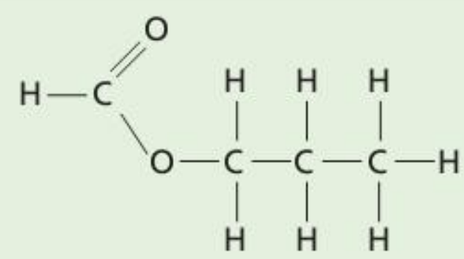
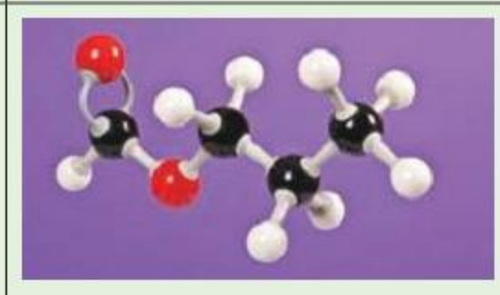
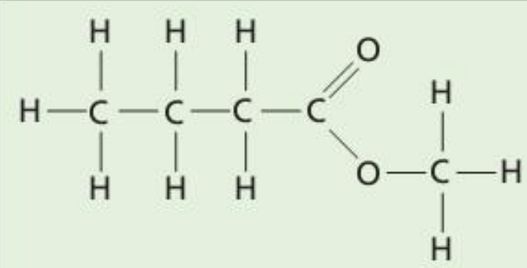
Ester	Made from		Structure	3D model
	Alcohol	Carboxylic acid		
Ethyl ethanoate $CH_3COOC_2H_5$	Ethanol C_2H_5OH	Ethanoic acid CH_3COOH		
Propyl methanoate $HCOOC_3H_7$	Propan 1-ol C_3H_7OH	Methanoic acid $HCOOH$		
Methyl butanoate $C_3H_7COOCH_3$	Methanol CH_3OH	Butanoic acid C_3H_7COOH		

Figure 15.10 The molecules of esters look like the 3D models in the photographs.

Figure 15.10 shows the actual arrangement of the atoms in some of the members of this family.

Many esters occur naturally and are responsible for the flavours in fruits and the smells of flowers. They are used, therefore, in some food flavourings and in perfumes (Figure 15.11).

Figure 15.11 Perfumes contain esters.

Fats and oils are naturally occurring esters which are used as energy storage compounds by plants and animals. They possess the same linkage as Terylene but have different units (see p. 241).

Questions

1 Write the structural formula for propanoic acid.
2 Write word and balanced chemical equations for the esterification of propanoic acid with ethanol.
3 Which of the following organic chemicals are carboxylic acids or alcohols?
Hexanoic acid, $C_6H_{13}COOH$
Butanol, C_4H_9OH
Octane, C_8H_{18}
Nonane, C_9H_{20}
Methanoic acid, $HCOOH$
Pentene, C_5H_{10}
Hexanol, $C_6H_{13}OH$
State why you have chosen your answers.

Other carboxylic acids

Aspirin

Aspirin (Figure 15.12) is one of the most frequently used painkillers in the world. It is also able to reduce inflammation and fever and a low dose taken on a daily basis over the age of 50 may prevent heart attacks. It is derived from another acid, salicylic acid, which can be obtained from willow bark. Salicylic acid has the same medicinal properties as aspirin and has been known since 1829. Salicylic acid, however, caused stomach bleeding. The conversion of salicylic acid to aspirin reduced these problems, but aspirin still has some adverse effects on the stomach if taken in excess.

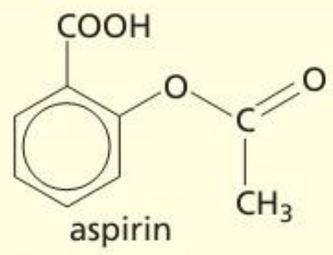


Figure 15.12 The structure of aspirin.

Ascorbic acid (vitamin C)

Vitamin C, also known as ascorbic acid, is an essential vitamin (Figure 15.13). Vitamin C is required by the body in very small amounts and it is obtained from foods.

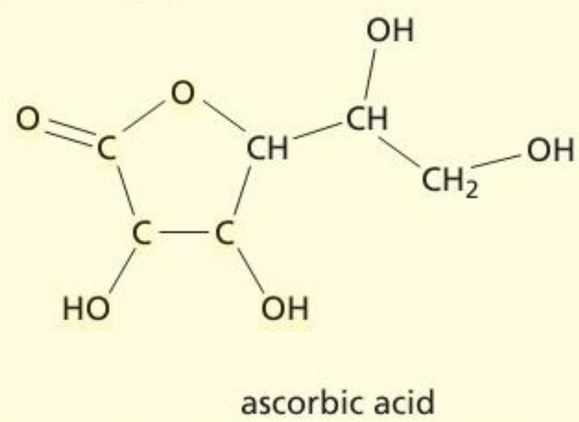


Figure 15.13 The structure of ascorbic acid (vitamin C).

Vitamin C is important to all animals, including humans, and is vital in the production of collagen. Collagen is important in the formation of connective tissues that give our body shape and help to support vital organs. Vitamin C prevents the disease scurvy. It is found in citrus fruits and brightly coloured vegetables, such as peppers and broccoli. Many people take vitamin C supplements, which are readily available from supermarkets and pharmacies (Figure 15.14). Although vitamin C is destroyed by exposure to air and heat, the average person usually reaches the recommended daily allowance of 60 mg through food.

Figure 15.14 These fruits and supplements all contain vitamin C.

Citric acid

Citric acid is an example of a tricarboxylic acid, one which contains three —COOH groups (Figure 15.15). It is an important acid and is found in all citrus fruits, for example lemons and oranges.

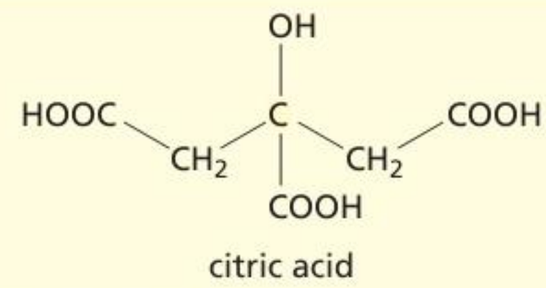


Figure 15.15 The structure of citric acid.

Soaps and detergents

Millions of tonnes of soaps and soapless detergents are manufactured worldwide every year. Soap is manufactured by heating natural fats and oils of either plants or animals with a strong alkali. These fats and oils, called triglycerides, are complicated ester molecules.

Fat is boiled with aqueous sodium hydroxide to form soap. The esters are broken down in the presence of water – hydrolysed. This type of reaction is called **saponification**. The equation given below is that for the saponification of glyceryl stearate (a fat).

glyceryl stearate + sodium hydroxide → sodium stearate (soap) + glycerol

$$\begin{array}{l} H_{35}C_{17}\text{–}COOCH_2 \\ \quad\quad\quad\quad | \\ H_{35}C_{17}\text{–}COOCH \\ \quad\quad\quad\quad | \\ H_{35}C_{17}\text{–}COOCH_2(l) \end{array} + 3NaOH(aq) \rightarrow 3C_{17}H_{35}COO^{-}Na^{+}(l) + \begin{array}{l} HO\text{–}CH_2 \\ \quad\quad | \\ HO\text{–}CH_2 \\ \quad\quad | \\ HO\text{–}CH_2(l) \end{array}$$

The cleaning properties of the soap depend on its structure and bonding. Sodium stearate consists of a long hydrocarbon chain which is hydrophobic (water hating) attached to an ionic 'head' which is hydrophilic (water loving) (Figure 15.16).

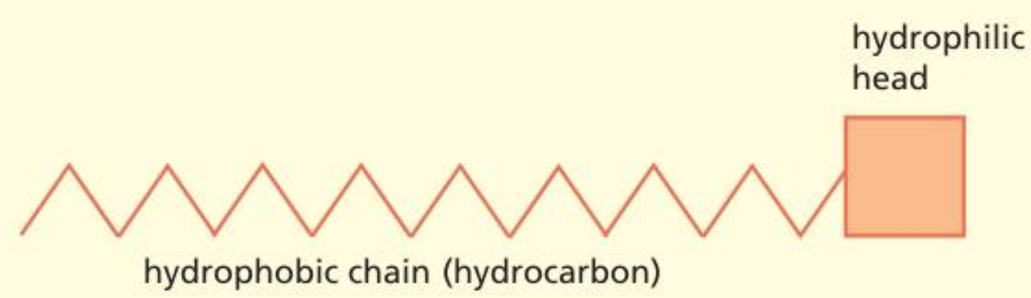


Figure 15.16 Simplified diagram of a soap molecule.

Covalent compounds are generally insoluble in water but they are more soluble in organic solvents. Ionic compounds are generally water soluble but tend to be insoluble in organic solvents. When a soap is put into water which has a greasy dish (or a greasy cloth) in it, the hydrophobic hydrocarbon chain on each soap molecule becomes attracted to the grease and becomes embedded in it (Figure 15.17).

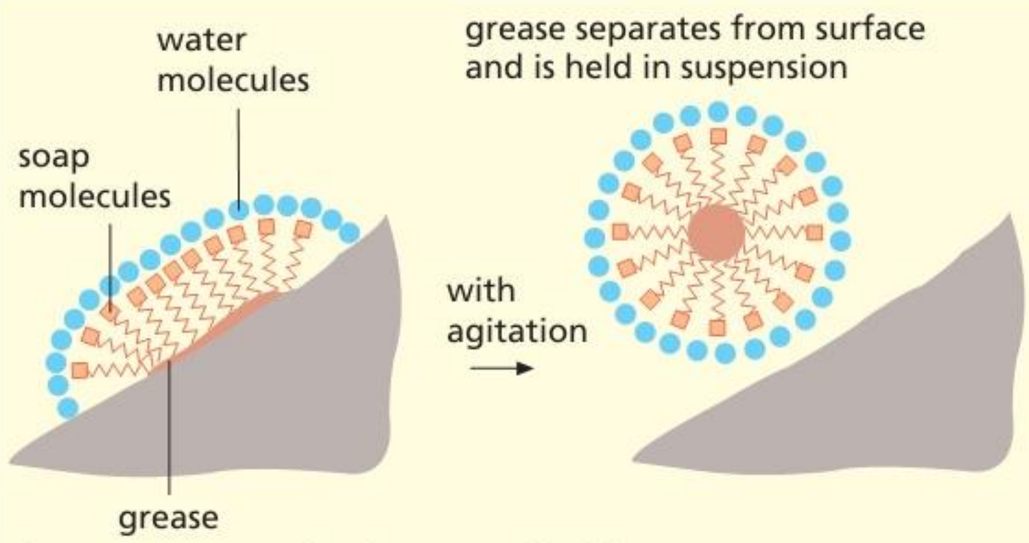


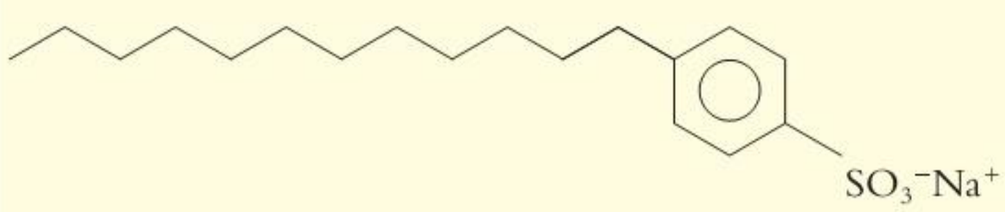
Figure 15.17 Soap dissolves grease like this.

On the other hand, the hydrophilic ionic head group is not attracted to the grease but is strongly attracted to the water molecules. When the water is stirred, the grease is slowly released and is completely surrounded by the soap molecules. The grease is, therefore, 'solubilised' and removed from the dish. The soap is able to remove the grease because of the combination of the covalent and ionic bonds present.

Soapless detergents

In Chapter 11, p. 188, we discussed the way in which, in hard water areas, an insoluble scum forms when soap is used. This problem has been overcome by the development of synthetic **soapless detergents**. These new substances do not form scum with hard water since they do not react with Ca^{2+} and Mg^{2+} present in such water. Furthermore, these new soapless detergent molecules have been designed so that they are biodegradable. Bacteria readily break down these new molecules so that they do not persist in the environment.

Sodium alkyl benzene sulfonates were developed in the early 1970s. The structure of sodium 3-dodecylbenzene sulfate, $C_{18}H_{29}SO_3Na$, is given below.



The calcium and magnesium salts of this detergent molecule are water soluble, so the problem of scum is solved. Very many of our washing powders (and liquids) contain this type of substance.

Manufacture of soapless detergents

Soapless detergents are detergents that can be used more effectively than soap, particularly in hard water areas (see Chapter 11, p. 188). They are fairly cheap to make and are rapidly replacing soaps.

The general process involves, initially, the reaction of a long, straight-chain alkene, such as dodecene ($CH_3(CH_2)_9CH{=}CH_2$), with benzene.

benzene	+	dodecene	→	dodecylbenzene
$C_6H_6(l)$	+	$CH_3(CH_2)_9CH{=}CH_2(l)$	→	$C_6H_5(CH_2)_{11}CH_3(l)$

The molecular formula of dodecylbenzene is $C_{18}H_{30}$. This compound is then reacted with concentrated sulfuric acid to give a compound that is known as a sulfonic acid.

dodecylbenzene	+	sulfuric acid	→	dodecylbenzene sulfonic acid	+	water
$C_{18}H_{30}(l)$	+	$H_2SO_4(l)$	→	$C_{18}H_{29}SO_3H(aq)$	+	$H_2O(l)$

Finally, this is reacted with the alkali sodium hydroxide, NaOH.

dodecylbenzene sulfonic acid	+	sodium hydroxide	→	sodium dodecylbenzene sulfonate	+	water
$C_{18}H_{29}SO_3H(aq)$	+	$NaOH(aq)$	→	$C_{18}H_{29}SO_3^-Na^+(aq)$ (soapless detergent)	+	$H_2O(l)$

These days, soapless detergents such as this are to be found in most washing powders and liquids.

Questions

1 What class of organic compound do substances like glyceryl stearate belong to?
2 What do you understand by the terms:
   a hydrophobic?
   b hydrophilic?
   c saponification?
3 What is the main advantage of detergents over soaps?

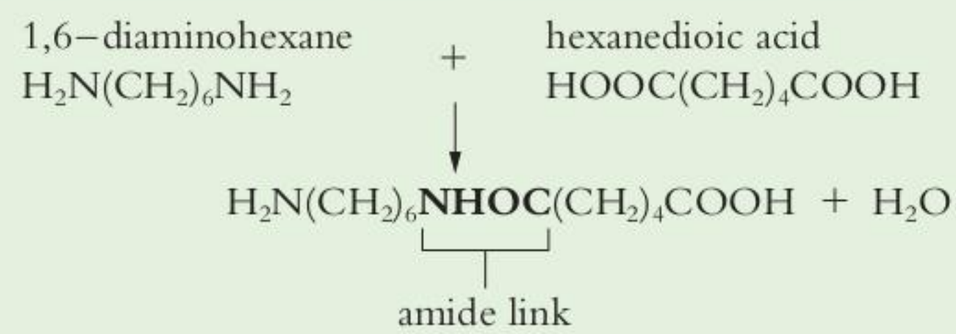
Condensation polymers

In Chapter 14 (p. 226) you studied the different addition polymers produced from alkenes. Not all polymers are formed by addition reactions, though. Some are produced as a result of a different type of reaction. In 1935 Wallace Carothers discovered a different sort of plastic when he developed the thermoplastic, nylon. Nylon is made by reacting two different chemicals together, unlike poly(ethene) which is made only from monomer units of ethene. Poly(ethene), formed by addition polymerisation, can be represented by:

–A–A–A–A–A–A–A–A–A–A–

where A = monomer.

The starting molecules for nylon are more complicated than those for poly(ethene) and are called 1,6-diaminohexane and hexanedioic acid.



The polymer chain is made up from the two starting molecules arranged alternately (Figure 15.18) as these molecules react and therefore link up. Each time a reaction takes place a molecule of water is lost.

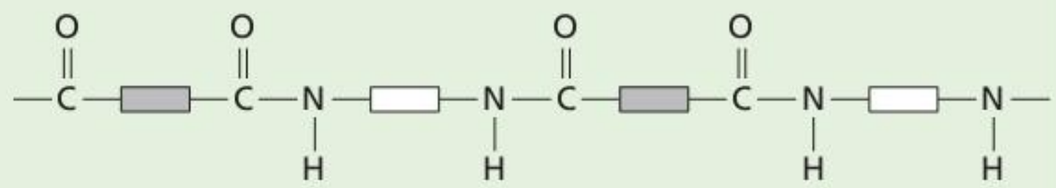
Figure 15.18 A nylon polymer chain is made up from the two molecules arranged alternately just like the two different coloured poppet beads in the photo.

This sort of reaction is called **condensation polymerisation**. This differs from addition polymerisation, where there is only one product. Because an amide link is formed during the polymerisation, nylon is known as a **polyamide**.

This is the same amide link as found in proteins (p. 243). It is often called the **peptide link**. This type of polymerisation, in which two kinds of monomer unit react, results in a chain of the type:

–A–B–A–B–A–B–A–B–A–B–

Generally, polyamides have the structure



When nylon is made in industry, it forms as a solid which is melted and forced through small holes (Figure 15.19). The long filaments cool and solid nylon fibres are produced which are stretched to align the polymer molecules and then dried. The resulting yarn can be woven into fabric to make shirts, ties, sheets and parachutes or turned into ropes or racket strings for tennis and badminton rackets. The annual worldwide production of nylon is expected to reach 6 million tonnes by 2015.

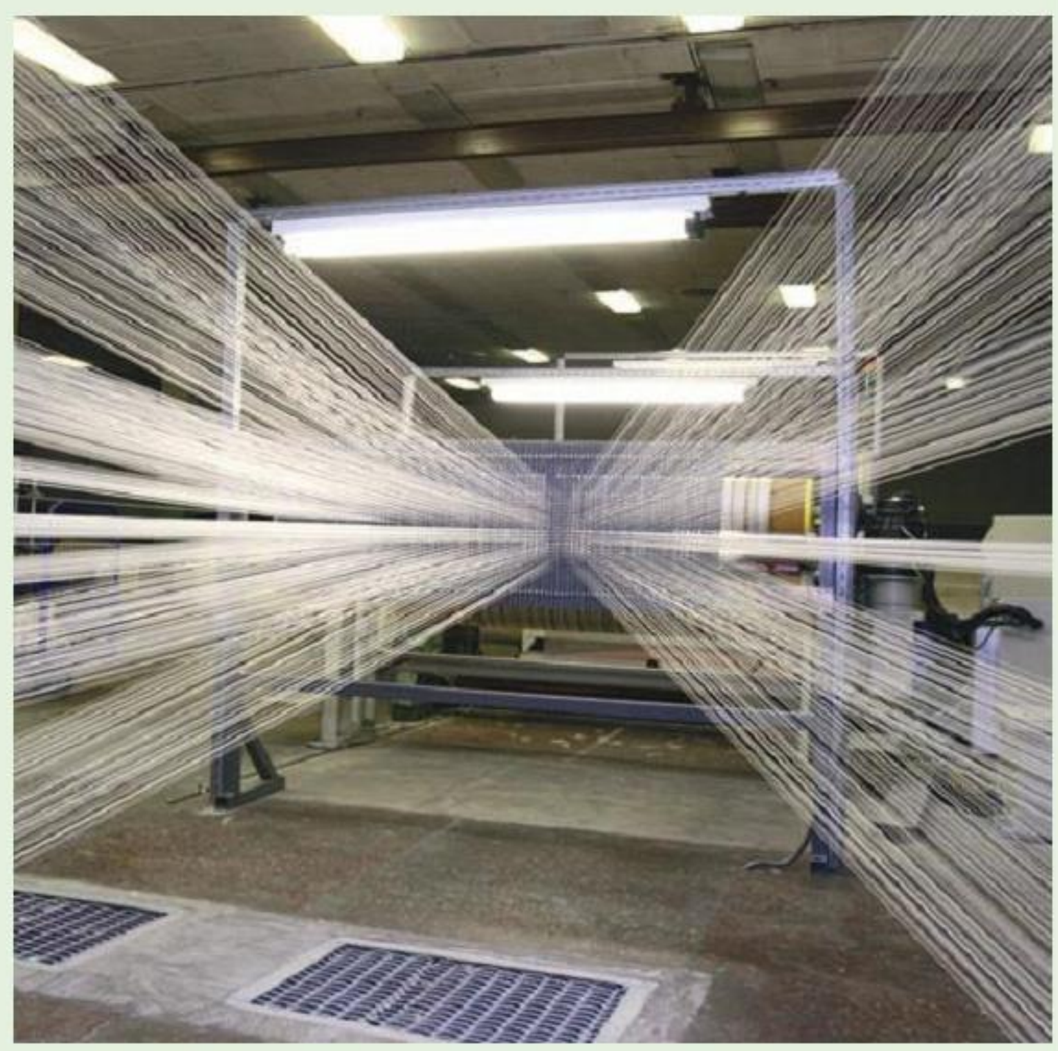
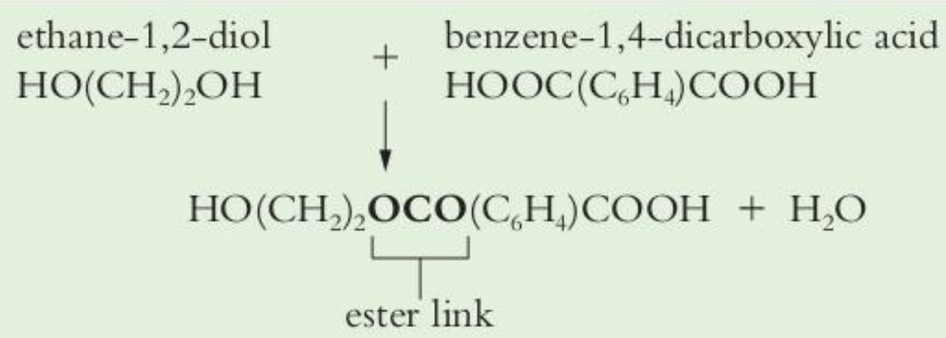
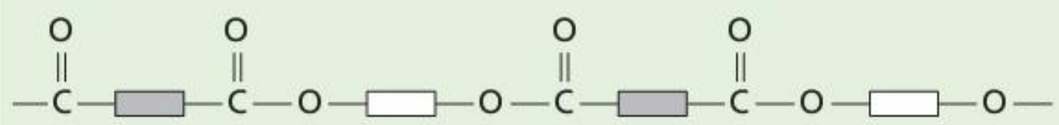


Figure 15.19 Nylon fibre is formed by forcing molten plastic through hundreds of tiny holes.

We can obtain different polymers with different properties if we carry out condensation polymerisation reactions between other monomer molecules. For example, if we react ethane-1,2-diol with benzene-1,4-dicarboxylic acid, then we produce a polymer called Terylene.



This ester link is the same linkage as in fats. Generally, polyesters have the structure



Like nylon, Terylene can be turned into yarn, which can then be woven. Terylene clothing is generally softer than that made from nylon but both are hard wearing. Because an ester link is formed during the polymerisation, Terylene is known as a **polyester**.

Fats possess the same sort of linkage as Terylene but have different units. For a further discussion see p. 238.

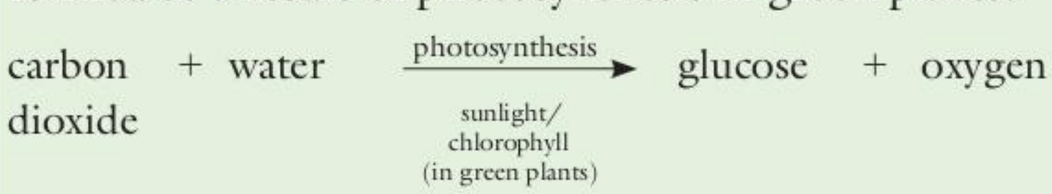
Questions

1 Draw the structure of the repeating units found in:
   a nylon b Terylene.
2 Explain the differences between an addition polymer and a condensation polymer.

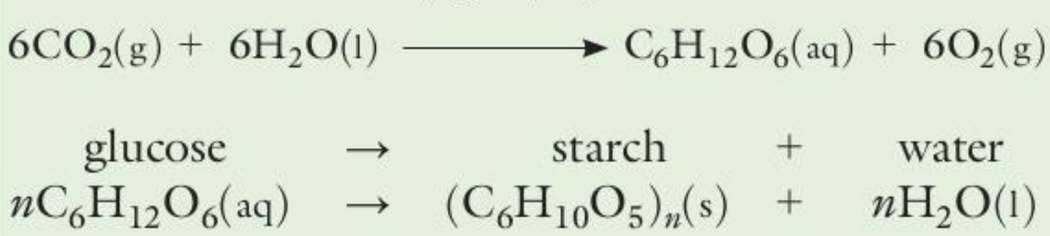
Some biopolymers

Starch

Starch is a **biopolymer** or **natural polymer**. It is a condensation polymer of glucose, a type of sugar. It is often produced as a way of storing energy and is formed as a result of photosynthesis in green plants.



$$6CO_2(g) + 6H_2O(l) \longrightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$



Starch can be represented by

Both starch and glucose are **carbohydrates**, a class of naturally occurring organic compounds which can be represented by the general formula $(CH_2O)x$.

Along with fats and proteins, they are among the main constituents of food.

Starch occurs in potatoes, rice and wheat. Glucose, from which starch is polymerised, belongs to a group of simple carbohydrates known as **monosaccharides**. They are sweet to taste and soluble in water. Starch belongs to the more complicated group of carbohydrates known as **polysaccharides**. Starch does not form a true solution and it does not have a sweet taste. With iodine it gives an intense blue colour (nearly black), which is used as a test for starch or iodine itself (Figure 15.20).

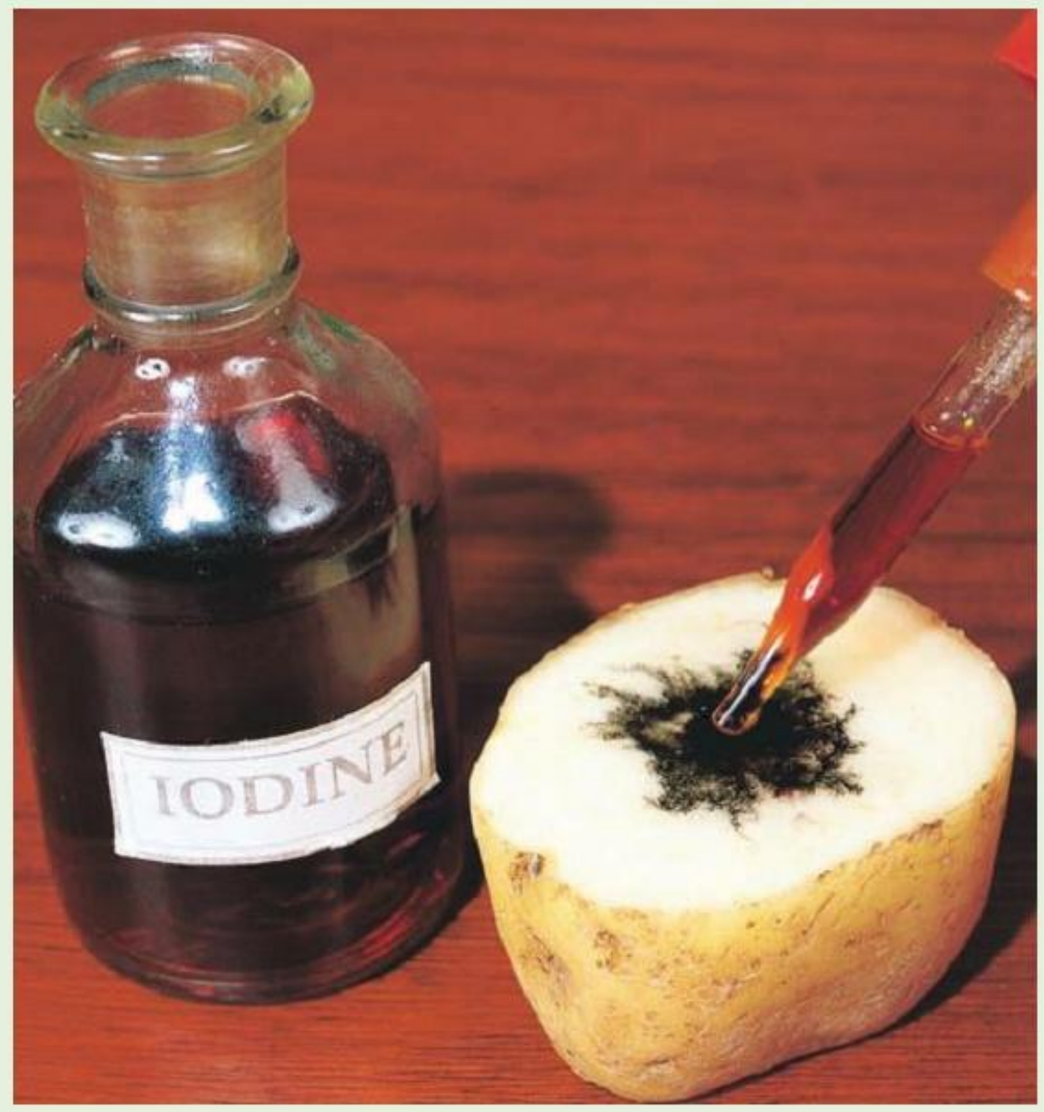
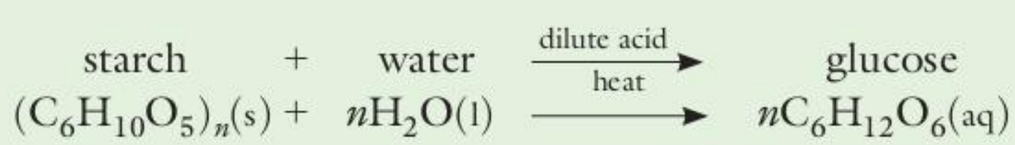


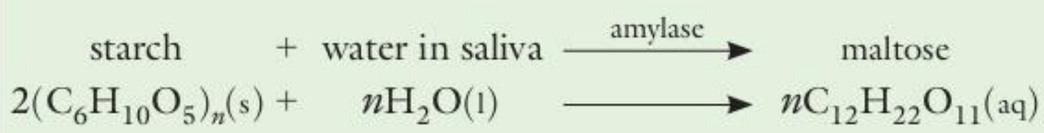
Figure 15.20 A dark blue–black colour is produced when dilute iodine solution is applied to starch, for example in a potato.

Hydrolysis of starch

Starch can be broken down in two ways, both of which take place in the presence of water. Hence the reactions are known as **hydrolysis** reactions. Hydrolysis of starch is the key reaction that enables us to use this energy source. If starch is boiled for about one hour with dilute hydrochloric acid, it is broken down into its monomers, glucose molecules.



If starch is mixed with saliva and left to stand for a few minutes, it will break down to maltose, a **disaccharide** (that is two joined monosaccharides). The enzyme present in the saliva, called amylase, catalyses this hydrolysis reaction.



Enzymes are very efficient natural catalysts present in plants and animals. They do not require high temperatures to break down the starch to maltose. In humans, a salivary amylase breaks down the starch in our food. If you chew on a piece of bread for several minutes, you will notice a sweet taste in your mouth. The above hydrolysis reactions are summarised in Figure 15.21.

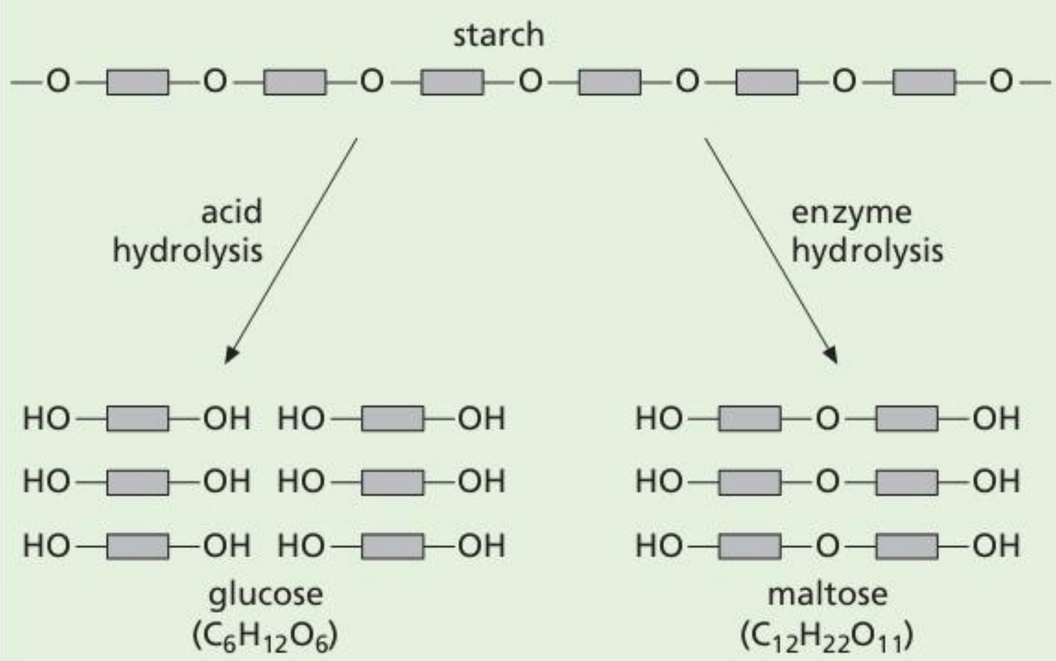


Figure 15.21 Starch produces glucose or maltose depending on the type of hydrolysis used. Using chromatography it is possible to separate and identify the products of the different styles of hydrolysis.

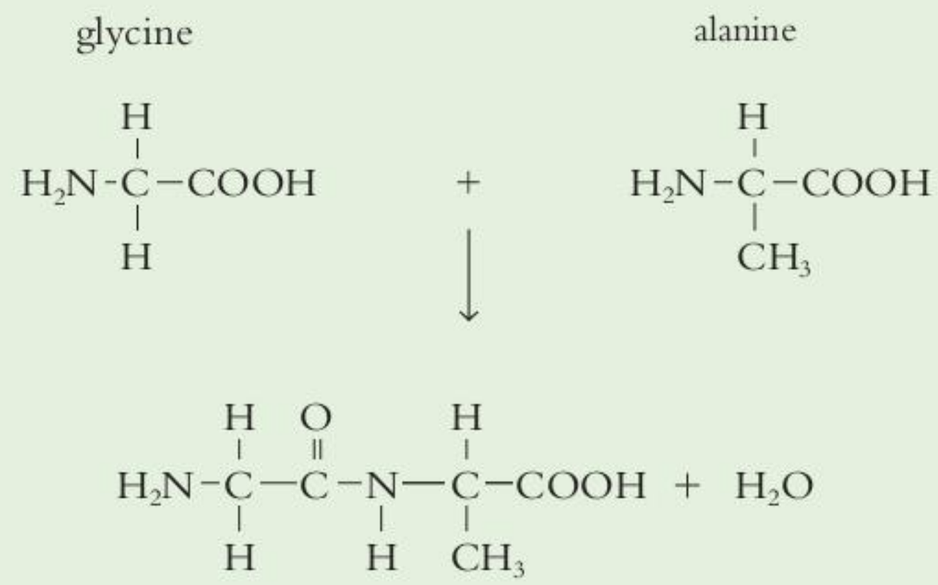
Questions

1. In the hydrolysis of starch, how, using a chemical test, could you tell whether all the starch had been hydrolysed?
2. Describe a method you could possibly use to identify the products of the different types of hydrolysis.

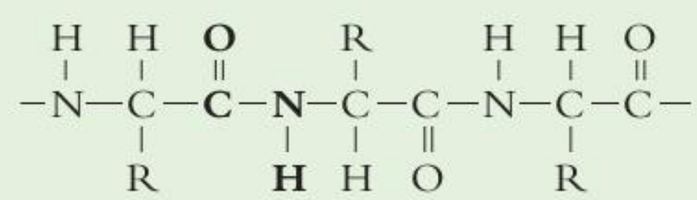
Amino acids

There are 20 different amino acids and they each possess two functional groups. One is the carboxylic acid group, –COOH. The other is the amine group, $–NH_2$. The two amino acids shown at the top of the next column are glycine and alanine.

Amino acids are the building blocks of **proteins**. Similar to nylon (see p. 241) proteins are polyamides, as they contain the –CONH– group, which is called the amide or, in the case of proteins, the peptide link. Proteins are formed by condensation polymerisation.



Protein chains formed by the reaction of many amino acid molecules have the general structure shown below.



Further reaction with many more amino acids takes place at each end of each molecule to produce the final protein (Figures 15.22 and 15.23). For a molecule to be a protein, there must be at least 100 amino acids involved. Below this number, they are called polypeptides. Proteins make up some 15% of our body weight.

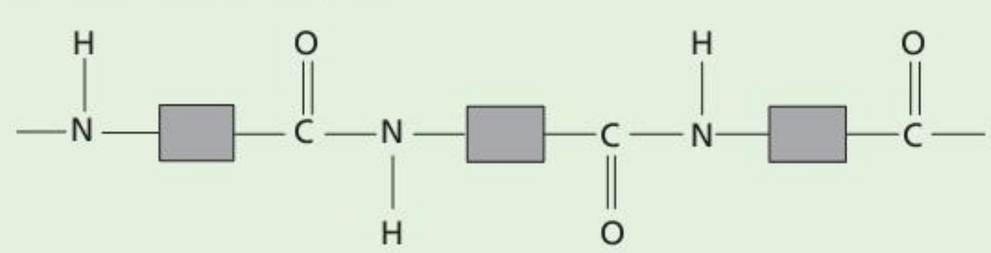


Figure 15.22 General structure of a protein.

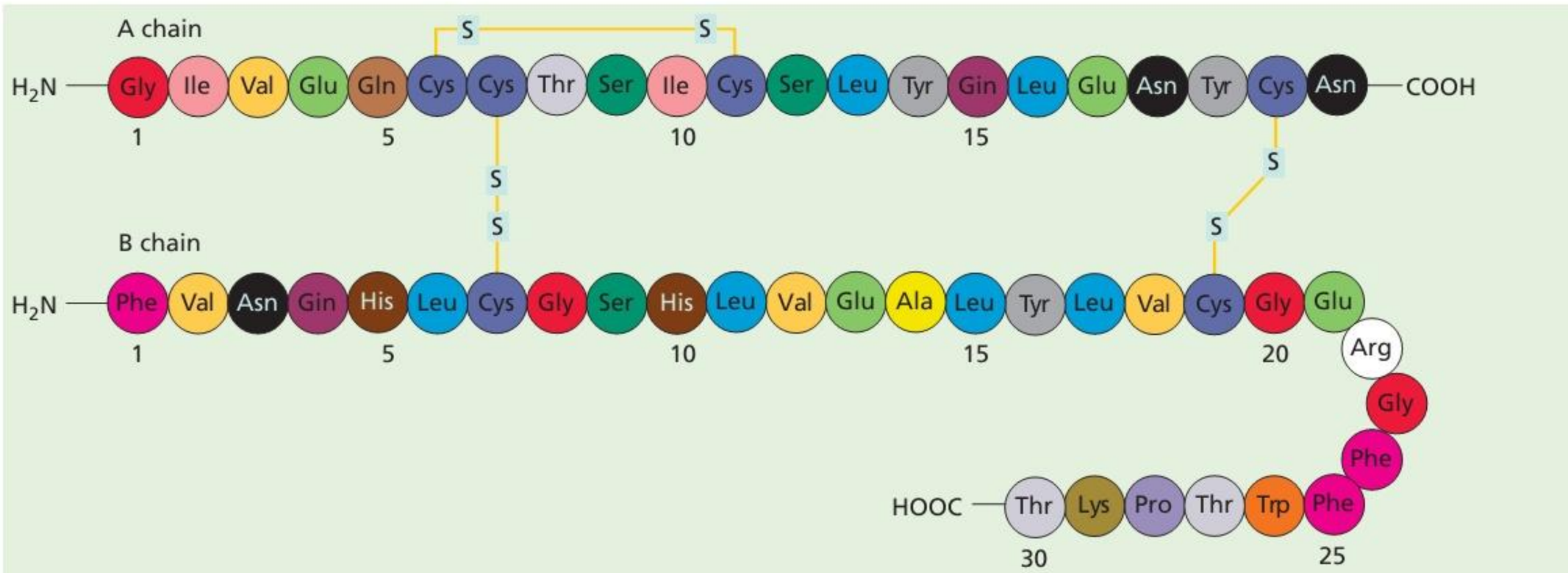


Figure 15.23 The structure of a protein – human insulin (the different coloured circles represent different amino acids in this protein).

Proteins fall broadly into two groups: they can be fibrous or globular.

- Fibrous proteins – these have linear molecules, are insoluble in water and resistant to alkalis and acids. Collagen (in tendons and muscles), keratin (in nails, hair, horn and feathers) and elastin (in arteries) are all fibrous proteins.
- Globular proteins – these have complicated three-dimensional structures and are soluble in water. They are easily affected by acids, alkalis and temperature increase, when they are said to be denatured. Casein (in milk), albumen (in egg white) and enzymes are examples of globular proteins.

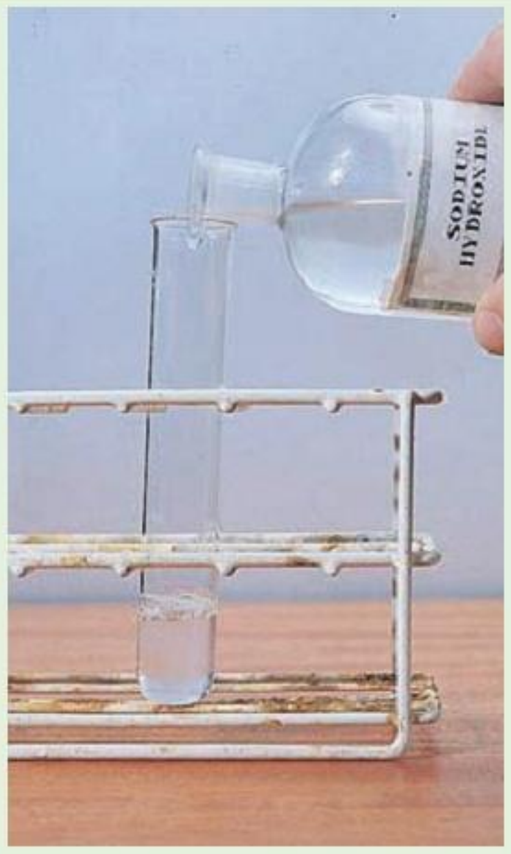
Analysis of the products of the hydrolysis of proteins

How can you determine which amino acids are present in a particular protein? This involves hydrolysis of the peptide (amide) bonds in the protein so that the individual amino acids are released. This can only be done by heating the protein with dilute hydrochloric acid. The mixture of amino acids is then separated by thin layer chromatography (TLC) (Figure 15.24) or electrophoresis. In both cases, a locating agent (Chapter 2, p. 23), such as ninhydrin, is used. This ensures that the spots of amino acid are visible.

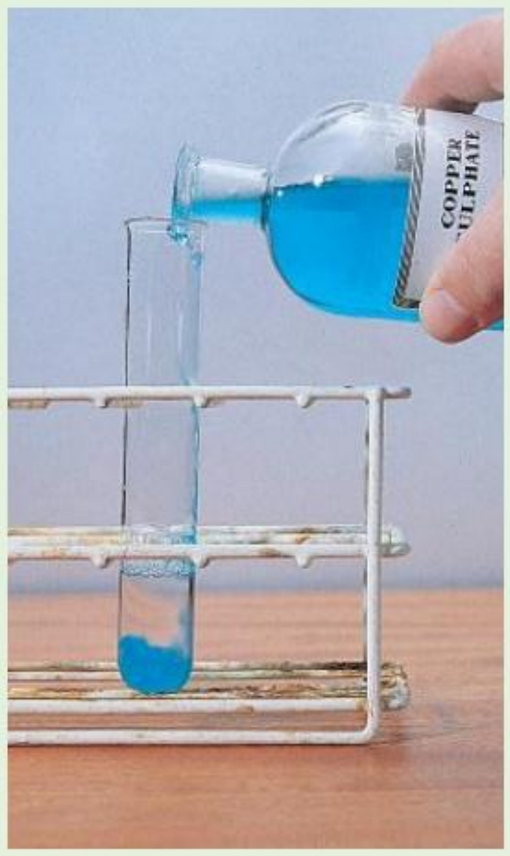
Figure 15.24 Amino acids can be separated and identified by TLC.

If you are trying to show only the presence of a protein, a quick test to carry out is known as the **Biuret test**. A mixture of dilute sodium hydroxide and 1% copper(II) sulfate solution is shaken with a sample of the material under test. If a protein is present, a purple colour appears after about three minutes (Figure 15.25).

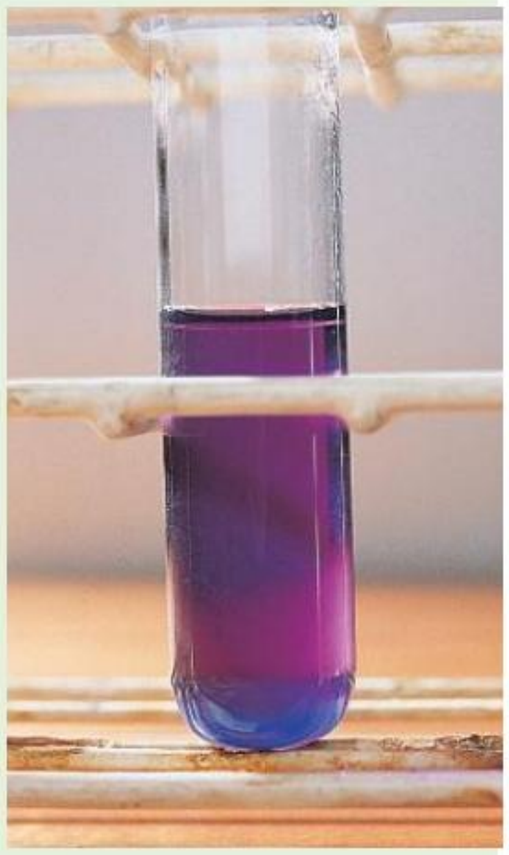
a Testing for a protein.



b Adding dilute sodium hydroxide.



c Adding 1% copper(II) sulfate.



d The purple colour shows the presence of a protein.

Figure 15.25 The Biuret test.

Figure 15.26 Crick and Watson based this model for DNA on X-ray studies performed by Rosalind Franklin and chemical analysis.

Genes are the units of heredity that control the characteristics of organisms. A gene is made of DNA. No two individuals have the same DNA sequence. DNA 'fingerprinting' has become a very powerful forensic science tool in the investigation of crime (Figure 15.27).

Figure 15.27 Identical patterns shown by DNA fingerprinting can identify the criminal.

DNA

Deoxyribonucleic acid (DNA) belongs to a group of chemicals called the nucleic acids (Figure 15.26). They are also biopolymers. DNA controls the prote synthesis within your cells. When you eat a food containing proteins, such as meat or cheese, your digestive enzymes break down the proteins present into individual amino acids. The DNA in your cells controls the order in which the amino acids are repolymerised to make the proteins you need!

Questions

1. Which two functional groups do amino acids possess?
2. How many amino acids have to be involved before the biopolymer is called a protein?
3. Name the process by which the individual amino acids in a protein are released by reaction with a dilute acid.
4. Explain how DNA fingerprinting may be used in paternity suits.

Pharmaceuticals

Pharmaceuticals are **drugs** that are prepared and sold with the intention of treating illness. A drug is any substance, natural or synthetic, which alters the way in which the body works. There are many categories of drugs. The following are some examples.

- Anaesthetics – these induce loss of feeling and/or consciousness, for example fluothane.
- Analgesics – these relieve pain, for example aspirin.
- Antibiotics – these are substances, for example penicillin, originally produced by micro-organisms, which are used to kill bacteria. However, most antibiotics are now made in chemical laboratories, for example carbenicillin.
- Sedatives – these induce sleep, for example barbiturates.
- Tranquillisers – these will give relief from anxiety, for example Valium.

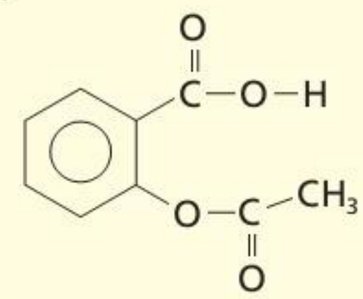
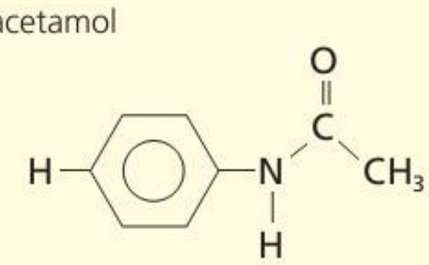
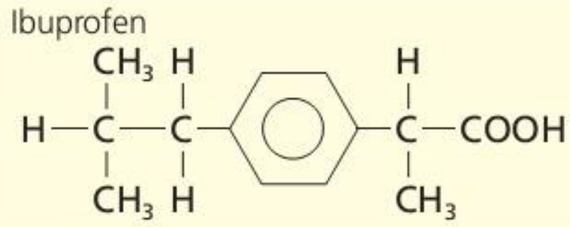
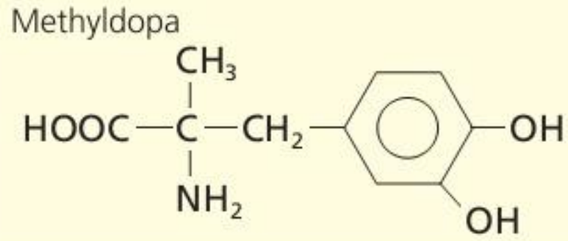
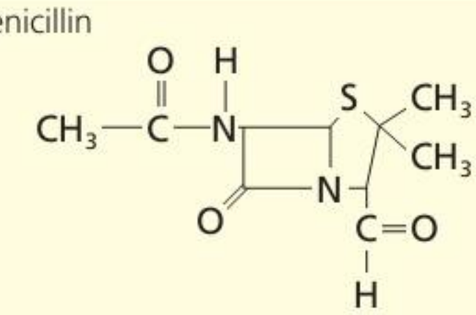
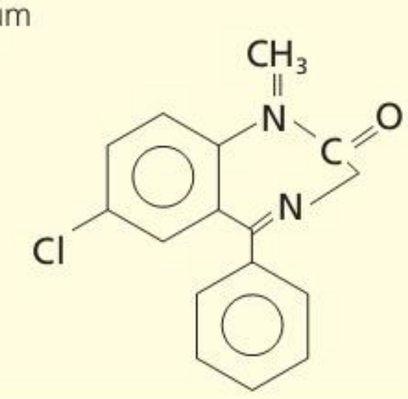
There are, of course, many other types of drug available which have very specific uses. For example, methyldopa was developed to relieve hypertension (high blood pressure), and antihistamines were developed to help control travel sickness, hayfever and allergic reactions.

The pharmaceutical industry is one of the most important parts of the chemical industry and is a major consumer of the products of the petrochemical industry. It is a high-profit industry but with very high research and development costs. For example, it costs in excess of £100 million to discover, test and get a single drug on to the market.

Today, the pharmaceutical industry could be called the 'medicines by design' industry. Companies such as GlaxoSmithKline have teams of chemists and biochemists working almost around the clock to discover, test, check for safety and produce drugs that can deal with almost every known illness.

Table 15.4 shows the structures of a selection of some of the more common drugs available at the present time, along with their uses. The common names for these drugs are used, since their systematic, theoretical names are extremely complex.

Table 15.4 Commonly available drugs.

Name and structure	For the treatment of
Aspirin 	Headaches, mild pain, heart disease
Paracetamol 	Headaches, mild pain
Ibuprofen 	Arthritis, fever, mild to moderate pain
Methyldopa 	High blood pressure (hypertension)
Penicillin 	A variety of bacterial infections
Valium 	Feelings of anxiety or depression

Drug abuse

Some of the very useful drugs developed by chemists can be habit forming. For example, barbiturates (in sleeping tablets) and amphetamines (stimulants) fall into this category. Another drug that has created problems in the past is Valium, which is not itself addictive but when used in the long term makes people dependent on it. Severe psychological and physiological problems can arise.

It should be noted, however, that it is the opiates which cause **addiction**. Cocaine and heroin are just two examples of such substances. Consequences of the addiction include personal neglect, both of nutritional needs and of hygiene. For a short-term feeling of well-being ('fix'), the addict is prepared to do almost anything. Addicts often turn to a life of crime to fulfil their cravings for the opiates. Addicts, especially those injecting drugs, are at a high risk of HIV (human immunodeficiency virus) infection as they often share needles with other drug addicts, who may be HIV positive. Public awareness campaigns aim to educate everyone on the dangers in society, including drug abuse and its related risks. (Figure 15.28)

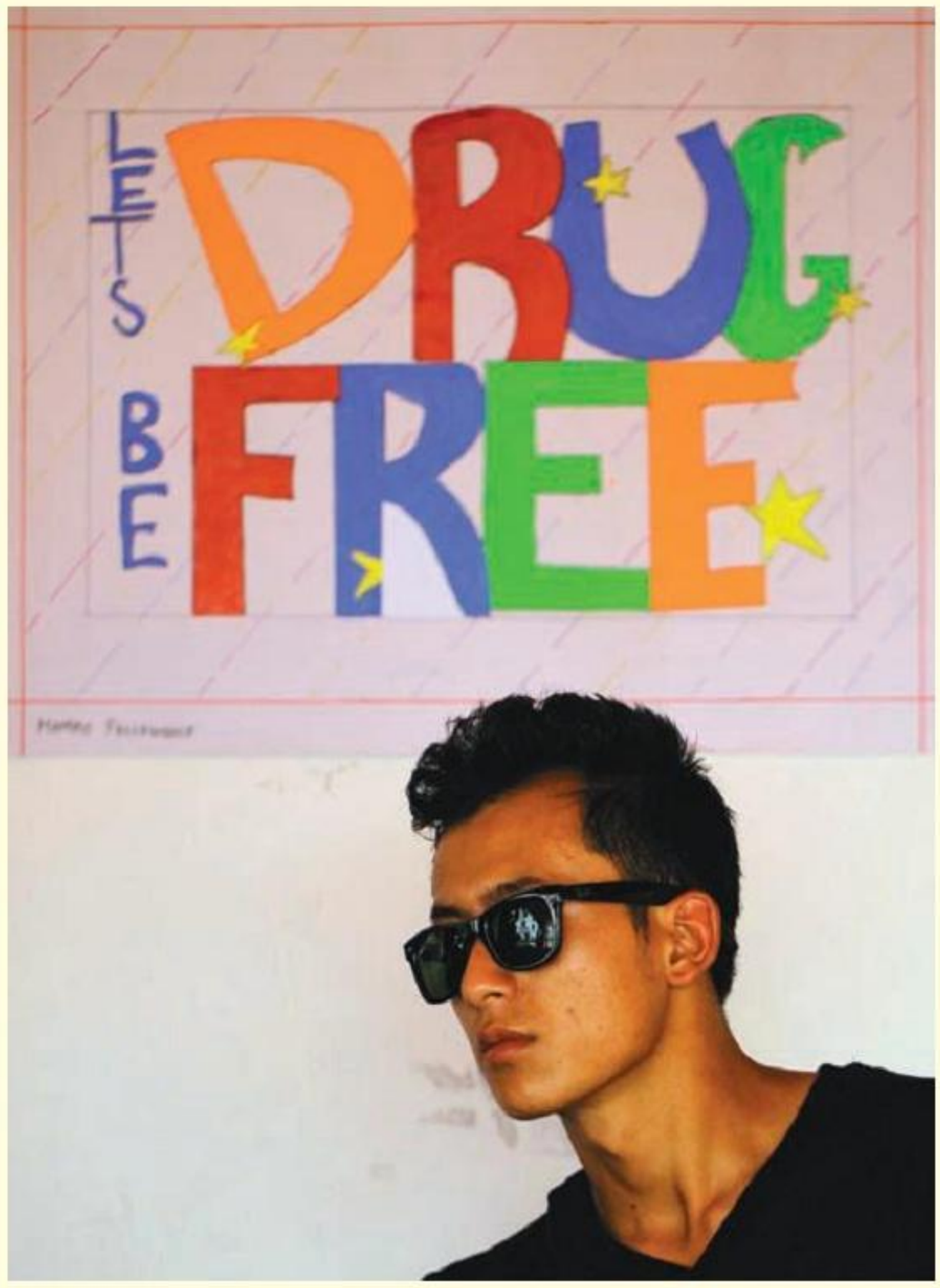


Figure 15.28 Increasing public awareness of the risks associated with drug abuse is very important.

Questions

1 Using the data given in Table 15.4, suggest which of the pharmaceuticals contains:
   a sulfur
   b an –OH group
   c an $–NH_2$ group
   d an ester group.
2 Drug abuse is a rapidly growing problem worldwide. Using the information given in this section as well as your research skills, make a list of the addictive drugs. Also explain the problems that drug abuse can cause.

Checklist

After studying Chapter 15 you should know and understand the following terms.

- **Aerobic respiration** Respiration that takes place in the presence of air.
- **Alcohols** Organic compounds containing the –OH group. They have the general formula $C_nH_{2n+1}OH$. Ethanol is by far the most important of the alcohols and is often just called 'alcohol'.
- **Amino acids** These naturally occurring organic compounds possess both an $–NH_2$ group and a –COOH group on adjacent carbon atoms. There are 20 naturally occurring amino acids, of which glycine is the simplest.
- **Anaerobic respiration** Respiration that takes place in the absence of air.
- **Biopolymers** Natural polymers such as starch and proteins.
- **Biotechnology** Making use of micro-organisms in industrial and commercial processes. For example, the process of fermentation is brought about by the enzymes in yeast.
- **Biuret test** The test for proteins. A mixture of dilute sodium hydroxide and 1% copper(II) sulfate solution is shaken with the material under test. A purple colour appears after about three minutes if a protein is present.
- **Carbohydrates** A group of naturally occurring organic compounds which can be represented by the general formula $(CH_2O)_x$.
- **Carboxylic acids** A family of organic compounds containing the functional group –COOH. They have the general formula $C_nH_{2n+1}COOH$. The most important and well known of these acids is ethanoic acid, which is the main constituent in vinegar. Ethanoic acid is produced by the oxidation of ethanol.
- **Condensation polymer** A polymer formed by a condensation reaction (one in which water is given out). For example, nylon is produced by the condensation reaction between 1,6-diaminohexane and hexanedioic acid.

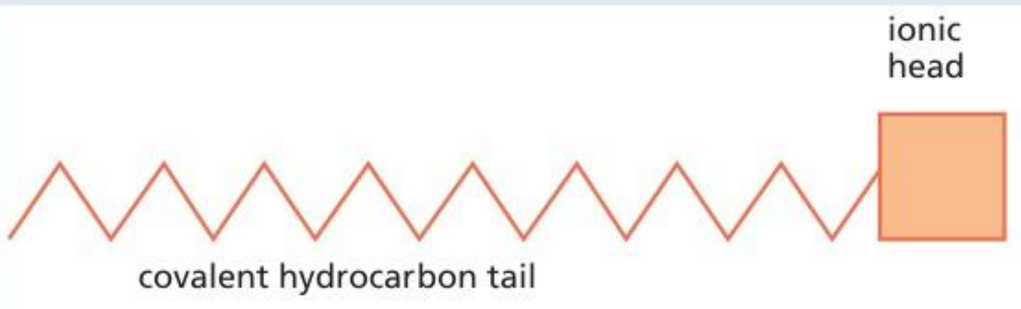
- **Dipeptide** Formed by the reaction of two amino acids.
- **Disaccharide** Two monosaccharides joined together, for example maltose.
- **DNA** Abbreviation for deoxyribonucleic acid. It belongs to a group of biopolymers called the nucleic acids. It is involved in the polymerisation of amino acids in a specific order to form the particular protein required by a cell.
- **Drug** Any substance, natural or synthetic, that alters the way in which the body works.
- **Drug abuse** This term usually applies to the misuse of addictive drugs, which include barbiturates and amphetamines, as well as the opiates, cocaine and heroin. These drugs create severe psychological and physiological problems. This leads to a variety of personal problems for the user.
- **Esters** A family of organic compounds formed by the reaction of an alcohol with a carboxylic acid in the presence of concentrated H_2SO_4. This type of reaction is known as esterification. Esters are characterised by a strong and pleasant smell (many occur in nature and account for the smell of flowers).
- **Fermentation** A series of biochemical reactions brought about by the enzymes in yeast or, more generally, by micro-organisms.
- **Functional group** The atom or group of atoms responsible for the characteristic reactions of a compound.
- **Genetic engineering** The deliberate manipulation of the genes in an organism.
- **HIV** Short for human immunodeficiency virus, from which AIDS (acquired immunodeficiency syndrome) can develop.
- **Hydrolysis** A chemical reaction involving the reaction of a compound with water. Acid hydrolysis usually involves dilute hydrochloric acid, and enzyme hydrolysis involves enzymes such as amylase.
- **Monosaccharides** A group of simple carbohydrates. They are sweet to taste and are water soluble (for example, glucose).
- **Pharmaceuticals** These are chemicals called drugs that are prepared and sold with the intention of treating disease (for example, methyldopa).
- **Polyamide** A condensation polymer, such as nylon, that contains the amide link, –NHOC–.
- **Polysaccharides** A group of more complicated carbohydrates. They generally do not form true solutions and do not have a sweet taste (for example, starch).
- **Proteins** Polymers of amino acids formed by condensation reactions. They fall broadly into two categories: fibrous proteins (for example, keratin and collagen) and globular proteins (for example, casein and albumen).
- **Saponification** A reaction in which an ester is heated with an alkali, such as sodium hydroxide, producing a free alcohol and an acid.
- **Soapless detergents** Soap-like molecules which do not form a scum with hard water. These substances have been developed from petrochemicals. Their calcium and magnesium salts are water soluble and they are biodegradable.
- **Soaps** Substances formed by saponification. In this reaction, the oil or fat (glyceryl ester) is hydrolysed by aqueous sodium hydroxide to produce the sodium salt of the fatty acid, particularly sodium stearate (from stearic acid). Soap will dissolve grease because of the dual nature of the soap molecule. It has a hydrophobic part (the hydrocarbon chain) and a hydrophilic part (the ionic head) and so will involve itself with both grease and water molecules. However, it forms a scum with hard water by reacting with the Ca^{2+} (or Mg^{2+}) present.

Organic chemistry 2

Additional questions

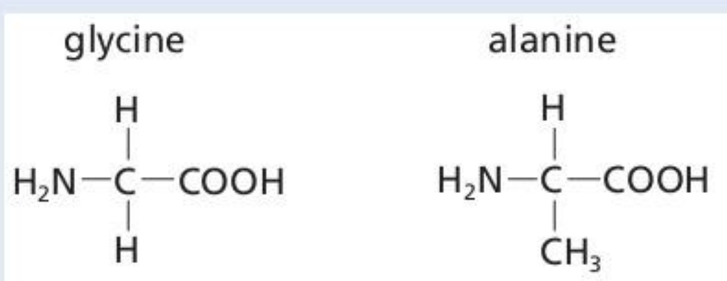
1 Explain the following.
   a The problem of plastic waste has been overcome.
   b The majority of detergents produced today are biodegradable.
   c In bread making, yeast is added to the mix and the dough left to stand for a period of time.
   d Poly(ethene) is a thermoplastic.

2 a A detergent molecule may be represented by the following simplified diagram.



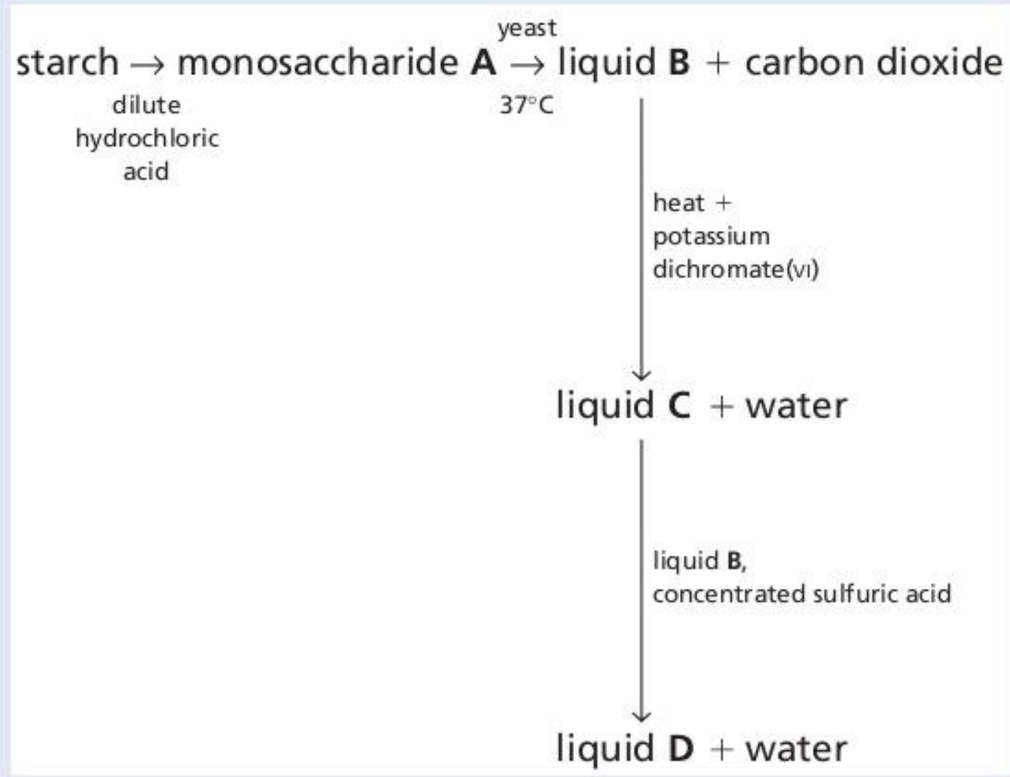
   Use this representation of a detergent molecule in a series of labelled diagrams to show how detergents can remove grease from a piece of greasy cloth.
   b Explain why detergents do not form a scum with hard water, whereas soaps do.
   c The modern detergents are biodegradable.
      (i) Explain what this statement means.
      (ii) Why is it necessary for detergents to be biodegradable?

3 A piece of cheese contains protein. Proteins are natural polymers made up of amino acids. There are 20 naturally occurring amino acids. The structures of two amino acids are shown below.



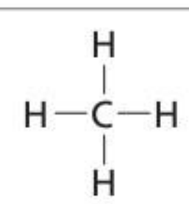
   a Name the type of polymerisation involved in protein formation.
   b Draw a structural formula to represent the part of the protein chain formed by the reaction between the amino acids shown above.
   c What is the name given to the common linkage present in protein molecules?
   d Why is there such a huge variety of proteins?
   e Name and describe the features of the two broad groups of proteins.

4



   a Name and give the formula and one use of each of the substances **A** to **D**.
   b Write word and balanced chemical equations for the reactions involved in the formation of liquids **B**, **C** and **D**.
   c Starch is classified as a natural polymer or 'biopolymer'. Explain the meaning of this statement.
   d Name the processes by which:
      (i) starch is broken down
      (ii) liquid **B** is formed
      (iii) liquid **C** is formed
      (iv) liquid **D** is formed.

5 a Copy the following table and complete it by writing the structural formulae for methanol and methanoic acid.

Methane	Methanol	Methanoic acid
		

   b Describe a simple chemical test that could be used to distinguish methanol from methanoic acid.
   c (i) Name the class of compound produced when methanol reacts with methanoic acid.
      (ii) Name the type of reaction taking place.
      (iii) Write a word and balanced chemical equation for this reaction.
      (iv) Give two uses related to the class of compound formed in this reaction.

d The following reaction takes place when methanol is burned:

$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g)$
$\Delta H = -1452\,kJ\,mol^{-1}$

(i) How much heat energy would be liberated by burning:
0.5 mol of methanol?
4.0 mol of methanol?
4 g of methanol?

(ii) Calculate the volume of carbon dioxide produced at room temperature and pressure (rtp) when 16 g of methanol are burned. (A_r: H = 1; C = 12; O = 16. One mole of any gas at rtp occupies 24 dm^3.)

6 Why is it safe for us to use vinegar, which contains ethanoic acid, on food while it would be extremely dangerous for us to use dilute nitric acid for the same purpose?

7 a Which carboxylic acids and alcohols would you use to make the following esters?
(i) butyl methanoate
(ii) ethyl methanoate
(iii) propyl propanoate

b Which esters would be made using the following carboxylic acids and alcohols?
(i) ethanoic acid and butan-1-ol
(ii) methanoic acid and methanol
(iii) propanoic acid and butan-1-ol

c Draw the structures of the following esters:
(i) ethyl methanoate
(ii) propyl ethanoate.

d Give two uses related to the class of compound formed in this reaction.

8 a Name the polymerisation process that is used to make both nylon and Terylene.

b Name the starting materials for making:
(i) nylon
(ii) Terylene.

c Give the name and formula of the small molecule produced during the polymerisation reactions to produce both nylon and Terylene.

d Give the name of the chemical link that holds together:
(i) nylon
(ii) Terylene.

e Give two uses for:
(i) nylon
(ii) Terylene.

f Explain the difference between the type of polymerisation you have named in part **a** and addition polymerisation.

9 What are the advantages and disadvantages of the two methods used to manufacture ethanol:
a by fermentation
b by the hydration of ethene with steam?

16 Experimental chemistry

Objectives for experimental skills and investigations

Suggestions for practical work and assessment
Safety
Redox

Notes on qualitative analysis
Appearance or smell
Flame colours
Tests for aqueous cations
Tests for aqueous anions
Tests for gases

Objectives for experimental skills and investigations

The IGCSE chemistry syllabus suggests that students should be able to:

1. know how to safely use techniques, apparatus and materials (including following a sequence of instructions where appropriate)
2. plan experiments and investigations
3. make and record observations, measurements and estimates
4. interpret and evaluate experimental observations and data
5. evaluate methods and suggest possible improvements.

All candidates must take a practical paper based on the criteria above. You must take one of the following papers:

- Paper 5 – Practical Test, or
- Paper 6 – Alternative to Practical (written paper).

Some examples of exam style questions from Paper 6 are given on pp. 264–74. The practical assessment is 20% of the available marks.

Suggestions for practical work and assessment

The following list suggests practical exercises that could be used to both support the assessment objectives given above and enrich the study of chemistry; in addition, it should suggest plenty of opportunities for the enhancement and assessment of practical skills. The list is neither exhaustive nor prescriptive: the actual selection of practical work, whether from this list or from elsewhere, has to be governed by local factors such as the facilities available, safety considerations and the ability of the class.

Safety

In the suggested practical exercises, materials are used which, although familiar in many cases, are of a potentially hazardous nature, and appropriate care and precautions should be taken. We believe that the experiments can be carried out safely in school laboratories. However, it is the responsibility of the teacher to make the final decision depending on the circumstances at the time. Goggles should be worn at all times. In certain cases disposable gloves and fume cupboards will be required.

Teachers must ensure that they follow the safety guidelines set down by their employers, and a risk assessment must be completed for *any* experiment that is carried out. Teachers should draw students' attention to the hazards involved in the particular exercise to be performed.

Redox

The concept of redox, either in its elementary form (that is, as the loss or gain of oxygen) or as electron transfer, is a constant theme throughout the syllabus and the associated practical work. The list suggests plenty of varied examples, designated by **R**, to reinforce the theory.

Chapter 1 The particulate nature of matter

Solids, liquids and gases

- A simple demonstration to illustrate the three states of matter and their interconversions – ice, water and steam.
- Demonstrate the compressibility of gases and the incompressibility of liquids by using a syringe to show that the volume of a gas (e.g. air) decreases with pressure whereas that of a liquid (e.g. water) does not.

Changes of state

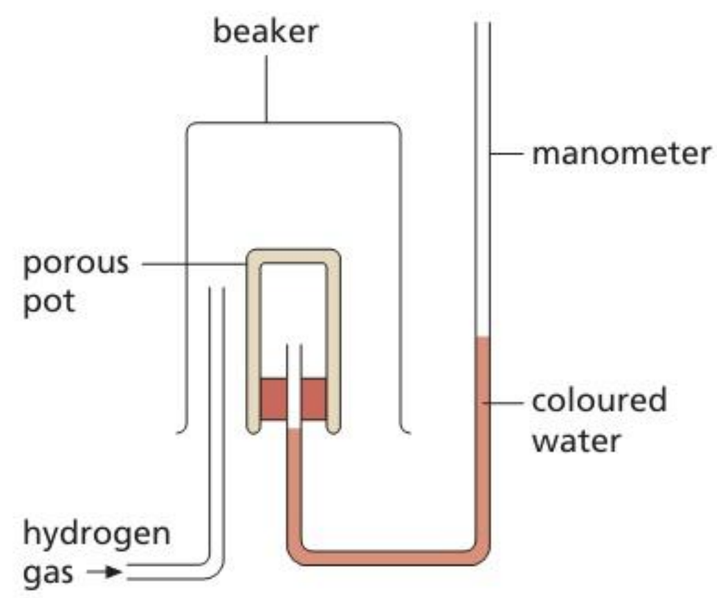
- Measure the melting points of both a pure sample and an impure sample of a solid; this will demonstrate that melting point is a means of assessing purity.
  - Measure the temperature of melting ice, then repeat after adding salt to the ice.
  - Measure the melting point of octadecanoic acid using a water bath (Figure 1.12, p. 6), then repeat using a sample that is contaminated with a trace of salol (phenyl 2-hydroxybenzoate).
  - Use a liquid paraffin bath or electrical heating apparatus to measure the melting point of an unknown compound and decide whether it is a pure sample.
  - Boiling point as a criterion of purity. Use a test-tube and thermometer to measure the boiling point of a suitable liquid. A simple example would involve heating calcium chloride solution in a test-tube with a few anti-bumping granules. Note that the bulb of the thermometer must be in the liquid itself to measure its boiling point (typically 102 °C in this case) and not in the vapour above the liquid (where it would register 100 °C).
- The fractional distillation of a mixture of water and ethanol (Figure 2.22, p. 21) illustrates that a pure liquid boils at one temperature but a mixture of liquids boils over a range of temperatures.
- A liquid will boil at a lower temperature when the external pressure is reduced, e.g. by using a water pump. The variation of boiling point with external pressure can be demonstrated with a volatile liquid, such as methanol; it is advisable to introduce some form of cold trap to reduce the amount of vapour entering the waste. An alternative is to allow boiling water to cool and so stop boiling, and then reduce the pressure: it will start to boil again at a temperature below 100 °C .
- Iodine, ammonium chloride or dry ice can be used to demonstrate sublimation. Heat the solid ammonium chloride or iodine in a test-tube to show that it disappears from the bottom of the tube and condenses higher up the tube, where it is cooler.

Heating and cooling curves

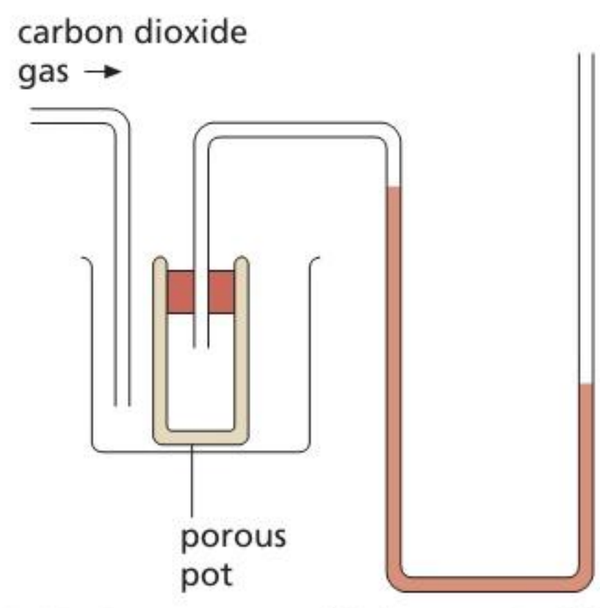
- Students will already be familiar with melting and boiling, so a more profitable approach might be to supply them with data to construct and interpret these curves.

Diffusion

- Brownian motion can be demonstrated using pollen or smoke. Smoke cells are commercially available.
- Illustrate gaseous diffusion using a gas jar of bromine vapour and a gas jar of air as described in the text (Figure 1.13, p. 6).
- An additional demonstration of gaseous diffusion is similar, but uses a gas jar of hydrogen *above* a gas jar of carbon dioxide. Both gas jars can be tested for the presence of carbon dioxide. It is preferable to set up this experiment twice, then, using the second set, hydrogen can be shown to be in both jars.
- Demonstrate the ammonia and hydrochloric acid experiment described in the text (Figure 1.14, p. 7). This also illustrates the dependence of the rate of diffusion on the relative molecular mass of a gas, M_r.
- An alternative is an arrangement of manometer and porous pot which effectively compares the rate of diffusion of a gas with that of air (Figure 16.1). This can be used to show that carbon dioxide diffuses slower than air and hydrogen diffuses faster.



a The pressure inside the porous pot rises because hydrogen diffuses in faster than air diffuses out.



b The pressure inside the porous pot falls because air diffuses out faster than carbon dioxide diffuses in.

Figure 16.1 Comparing the rate of diffusion of gases with that of air.

- Diffusion of a coloured solute in water. A large crystal of copper(II) sulfate is added to a measuring cylinder of water and is observed over several days.

Chapter 2 Elements, compounds and experimental techniques

Elements

- Display samples of as many elements as possible, either on information cards or on a copy of the Periodic Table. This is an effective observation exercise.
- Set up a circus of activities which includes both collecting data from computers or data books (e.g. melting and boiling points and density) and practical exercises on comparing electrical conductivities. These activities will illustrate the physical differences between metals and non-metals.

Compounds and mixtures

- Demonstrate the combination of elements to form compounds, for example magnesium and oxygen, iron and chlorine, hydrogen and oxygen, zinc and sulfur, carbon and oxygen, and aluminium and iodine.
- Demonstrate the iron and sulfur experiment described in the text (p. 16).

Separating mixtures

- Filtration, decanting and centrifuging to separate a solid from a liquid (p. 18).
- Evaporation to obtain a solute from a solution (p. 19).
- Crystallisation of impure benzoic acid. Benzoic acid is contaminated with a *trace* of a blue dye and then dissolved in the minimum amount of boiling water. Allow to cool and separate crystals by filtration. The blue colour will have disappeared or become fainter, showing that crystallisation has removed the impurity.
- Simple distillation to separate a solvent from a solution (p. 20).
- Separation of immiscible liquids using a separating funnel (p. 20).
- Fractional distillation of a mixture of ethanol and water will separate two miscible liquids and illustrates that an impure liquid does not boil at a specific temperature but over a range of temperatures (p. 21).
- Illustrate the use of chromatography as a test of purity and an analytical tool. Suitable examples include coloured inks, food colouring, pigments from flowers or grass, metallic cations and identification of sugars. At least one exercise should involve developing a chromatogram and at least one should involve the measurement of R_f values.

Gels, sols, foams and emulsions

- Set up a circus of gels, sols, emulsions and foams, e.g. 'jelly', milk, pumice stone, polyurethane foam, bread, emulsion paint, cola, hair cream, aerosol dispenser, salad cream. A silica gel can be made from sodium silicate and hydrochloric acid. Classify the examples according to type of colloidal system.
- Make a colloidal sol of iron(III) hydroxide by adding aqueous iron(III) chloride to boiling water, or a colloidal sulfur sol by adding dilute aqueous sodium thiosulfate to hydrochloric acid. For both of these sols, it can be shown that the solid phase is not separated by filtration.

Chapter 3 Atomic structure and bonding

Ionic bonding

- Look up the melting and boiling points of ionic compounds in data books.
- Show that most ionic compounds dissolve in water.
- Demonstrate that an ionic compound, potassium iodide or lithium chloride, conducts electricity in the molten state but not as a solid.

Covalent bonding

- Look up the melting and boiling points of simple covalent compounds in data books.
- Show that they are not good conductors of electricity in the liquid phase – demonstrate this with hexane, ethanol or phosphorus trichloride.
- Show that graphite is both soft and a good conductor of electricity.
- Use a glass cutter to illustrate the hardness of diamond.

Glasses and ceramics

- Heat a piece of glass tubing to demonstrate that it gradually softens, will bend and can be pulled into a finer tube. These changes show that glass is a supercooled liquid, not a solid.

Metallic bonding

- The properties of typical metals will be covered by the practical work on the transition elements in Chapter 9.

Chapter 4 Stoichiometry – chemical calculations

Moles and gases

- Measure the molar volume at rtp for hydrogen. Use a flask with a delivery tube and collect the gas over water in an inverted measuring cylinder or burette (Figure 16.2). Put an excess of dilute hydrochloric acid, say 100 cm^3 of 2 mol dm^{-3} acid, in the flask and add a piece of magnesium ribbon whose mass is known accurately and is about 0.1 g. (It is useful to measure the mass of a 1 m length of ribbon and then the mass of a smaller length can be calculated.) Measure the volume of hydrogen collected in the measuring cylinder or burette.

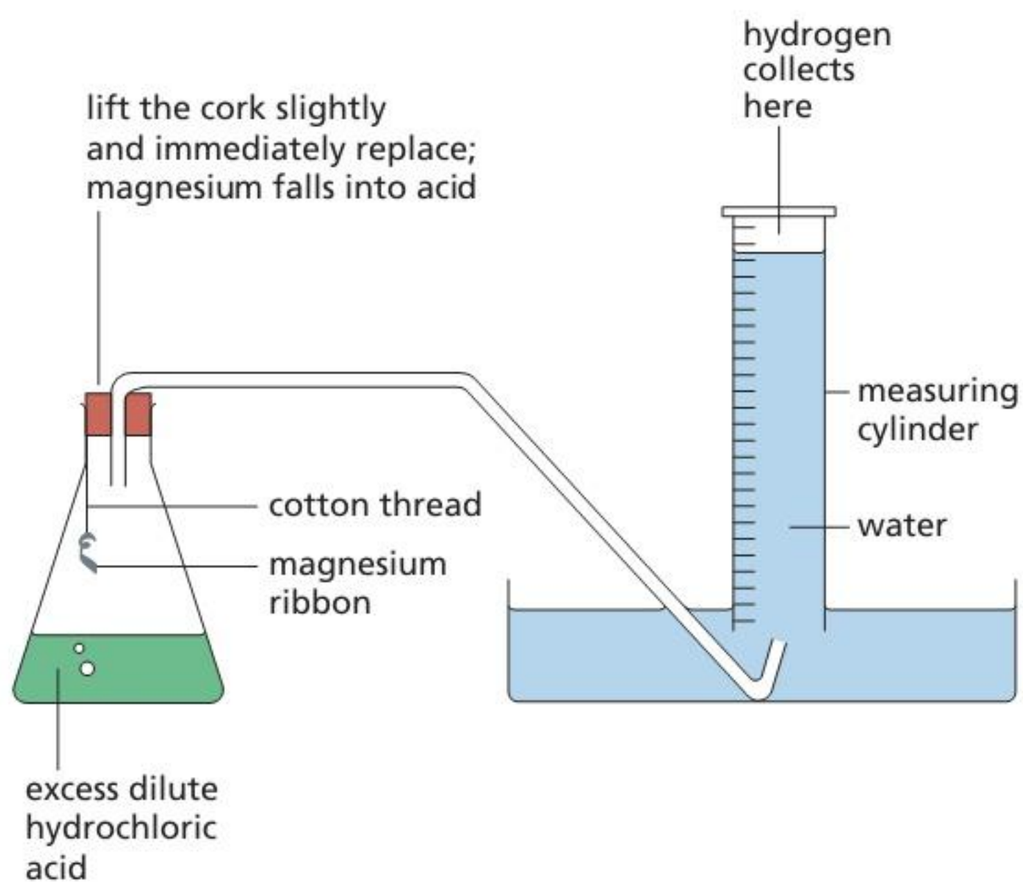


Figure 16.2 Measuring the molar volume of a gas at rtp.

- Measure the molar volume at rtp for carbon dioxide. Repeat the experiment above, using a known mass of a carbonate and collecting the carbon dioxide in a syringe.

Chapter 5 Electricity and chemistry

- Introductory practical. Electrolytes only conduct in the liquid phase, either molten or in aqueous solution. Experiments can be carried out to illustrate the differences between strong, weak and non-electrolytes. Use a circuit similar to that in Figure 5.2 (p. 72), but with an ammeter in preference to the lamp, and measure the conductivity of a selection of liquids, for example sodium chloride solution, hydrochloric acid, sodium hydroxide solution, water, dilute ethanoic acid, aqueous ammonia, paraffin and ethanol.
- Electrolysis as decomposition, for example of copper chloride solution, zinc bromide solution, hydrochloric acid and acidified water. Include tests for chlorine, oxygen and hydrogen. **R**
- Electrolysis of aqueous solutions of the following using inert electrodes: sodium chloride, copper(II) sulfate, sodium sulfate and sodium hydroxide.
- Electrolysis of aqueous copper(II) sulfate using copper electrodes.
- Examples of electroplating, such as nickel plating. Use a nickel anode, a copper cathode and nickel sulfate solution as the electrolyte in a cell similar to that in Figure 5.20 (p. 84).

Chapter 6 Chemical energetics

- Simple distillation of a synthetically manufactured 'crude oil' obtained from chemical suppliers.
- Inspection of some of the fractions of the primary distillation of crude oil – colour, viscosity, boiling point and flammability.
- Show examples of exothermic and endothermic reactions.
  - Dissolve ammonium nitrate in water – an example of an endothermic change.
  - Dissolve anhydrous copper(II) sulfate in water – an example of an exothermic change.
- Determination of the enthalpy of combustion of ethanol, as described in question 6 on p. 102.
- Show that the reaction between zinc and copper(II) ions is exothermic.
  - Add zinc dust to copper(II) sulfate solution. Note the temperature of the solution before and after the addition of the zinc. Here chemical energy is transformed into heat energy.
  - Set up a zinc/copper cell and note its polarity and voltage (Figure 6.23, p. 98). Here chemical energy is changed into electrical energy. Repeat using other cells, for example Cu/Ag, Fe/Cu and Mg/Cu. This will demonstrate that the reactivity series can be established using cells. **R**
- To simulate a fuel cell, electrolyse water for a few minutes until both electrodes are covered with bubbles of gas. Replace the power supply with a voltmeter. This will show that the cell produces a

voltage when the hydrogen on one electrode reacts with the oxygen on the other by the transfer of electrons through the external circuit. **R**

Chapter 7 Chemical reactions

- Reaction of calcium carbonate and acid (Figure 7.3, p. 105) – the effect of concentration and surface area.
- Reaction of sodium thiosulfate and acid (Figure 7.8, p. 107) – the effect of concentration.
- Decomposition of hydrogen peroxide, catalysed by manganese(IV) oxide – the effect of concentration and mass of catalyst on the rate and on the volume of product.

  First use a mixture of 2 cm^3 of 20 vol hydrogen peroxide and 48 cm^3 of water, with 0.2 g of manganese(IV) oxide. Then use a mixture of 4 cm^3 of 20 vol hydrogen peroxide and 46 cm^3 of water with 0.2 g of manganese dioxide. Finally use a mixture of 2 cm^3 of 20 vol hydrogen peroxide and 48 cm^3 of water with 0.4 g of manganese dioxide. Collect the oxygen over water in an inverted measuring cylinder or burette and obtain sets of readings of volume against time.

- Add a 2 cm length of magnesium ribbon to an excess of 2 mol dm^{-3} hydrochloric acid and measure the time taken for the metal to react. Repeat, with the same sized piece of magnesium but a different concentration of acid. (Another version of this experiment is to make temperature the variable and use hydrochloric acid at different temperatures.)
- Add a 2 cm length of magnesium ribbon to an excess of 2 mol dm^{-3} ethanoic acid. Measure the time taken for the metal to react and compare this result with the one above.
- Reaction of metal and acid – the effect of concentration and surface area on the reaction rate and on the volume of product.

  Collect the hydrogen and obtain data on the volume evolved against time. The apparatus used could be a flask with delivery tube and a gas syringe or burette. A length of magnesium ribbon is suspended above excess acid and is allowed to fall into the acid as the timing starts.

  The experiment could be repeated using different lengths of ribbon and different concentrations of acid. Also, a mass of magnesium powder equal to that of the ribbon could be suspended above the acid in a small tube.

- Decomposition of sodium chlorate(I), catalysed by a cobalt(II) salt.

  $$2NaClO(aq) \rightarrow 2NaCl(aq) + O_2(g)$$

  Use the same apparatus as in the previous exercise to obtain data on the volume of oxygen evolved against time. This can be analysed to investigate the effect of concentration on the reaction rate and on the volume of product.

- Identical mixtures of sodium thiosulfate solution and hydrochloric acid are used at different temperatures in the experiment described on p. 107. The results illustrate how rate depends on temperature when the concentrations are kept constant.
- Compare a silver halide that has been exposed to light with a sample kept in the dark. Prepare a suspension of silver bromide and divide it into two portions. Wrap one portion in aluminium foil and hold the other near a bright light.
- Divide a mixture of bromine and hexane into two portions. Expose one to bright light while the other is kept in the dark. Only the sample exposed to the light will evolve hydrogen bromide, and its brown colour will fade.
- Biology-type experiments on green leaves and light. It is probable that the students will already be familiar with these.
- The effect of enzymes on the decomposition of hydrogen peroxide. Add a piece of raw liver to 20 vol hydrogen peroxide.
- Add a few drops of washing-up liquid to 20 vol hydrogen peroxide. Divide this mixture into two portions and add manganese(IV) oxide to one portion. Compare the rate of decomposition of hydrogen peroxide in the two samples.

Chapter 8 Acids, bases and salts

Acids and alkalis

- With simple indicators, find the acid colour, the alkaline colour and the neutral colour. Hold the acid tube and the alkaline tube up to a light together and look through both to see the neutral colour.
- Dilution experiments using universal indicator and 0.1 mol dm^{-3} solutions of hydrochloric acid, sodium hydroxide, ethanoic acid and ammonia. These will establish the pH scale and promote an understanding of the distinction between strong and weak electrolytes.

- Demonstrate that a solution of hydrogen chloride in dry hexane does not exhibit acid properties. It does not conduct electricity, nor does it react with magnesium.

Formation of salts

- Reactions of a typical acid with metals. Use hydrochloric and/or sulfuric acid with magnesium, aluminium, zinc, iron and copper. **R**
- Reactions of typical acids with bases and carbonates. Use a selection of alkalis, insoluble bases and carbonates with hydrochloric, sulfuric and nitric acids.
- Small-scale preparation of a salt from a metal – magnesium sulfate-7-water (p. 123).
- Preparation of a salt by titration – sodium chloride or sodium sulfate-10-water (p. 124).
- Preparation of a salt from an insoluble carbonate and an acid – nickel sulfate-7-water (p. 123).
- Preparation of a salt from an acid and an insoluble base – copper(II) sulfate-5-water (p. 124).
- Tests for anions – nitrate, chloride, bromide, iodide, carbonate and sulfate (p. 262).
- Preparation of a number of insoluble salts by precipitation on a test-tube scale – an exercise in practical skills, observation, recording and writing equations.
- One formal preparation of an insoluble salt, such as lead iodide, to include filtering, washing and drying (p. 125).

Crystal hydrates

- Thermal decomposition of a hydrated salt – copper(II) sulfate-5-water. Add water carefully to the anhydrous salt (Figure 8.21, p. 128).
- Find x in $MgSO_4.xH_2O$ or $BaCl_2.xH_2O$ by thermal decomposition. Weigh a clean, dry crucible and lid. Add between 2 g and 3 g of the hydrated compound and re-weigh, then heat to constant mass. From the results calculate x.

Titration

- By titration find the concentration of a solution of an alkali using a standard acid, hydrochloric acid, as described in the text (p. 129).
- Using the same method (but a different calculation), find the concentration of a solution of an alkali using sulfuric acid.
- By titration find the mass of one mole of an acid, H_2Y, given a solution of H_2Y of known concentration in $g\,dm^{-3}$. A suitable acid would be ethanedioic acid, with phenolphthalein as the indicator.

Chapter 9 The Periodic Table

Group I – the alkali metals

- Teachers can demonstrate that lithium, sodium and potassium are soft metals and good conductors of electricity, and students can find the densities and melting points of these metals using data books.
- The metals are burnt in air or oxygen and then it is shown that their oxides are alkaline. Using a deflagrating spoon, the hot metal is placed into a gas jar of oxygen, the colour of the flame is noted and the pH of the white powder formed is measured.
- The reactions of lithium, sodium and potassium with cold water can be demonstrated to show that alkaline solutions are formed and to illustrate the difference in their reactivities.
- Sodium is burnt in chlorine to demonstrate the formation of sodium chloride.

Group II – the alkaline earth metals

- Similar exercises as for Group I can be conducted with magnesium and calcium. In addition, it can be demonstrated that magnesium reacts slowly with cold water but will burn in steam (Figure 10.3, p. 151).

Group VII – the halogens

- Investigate their physical properties by inspection and using data books.
- Demonstrate their displacement reactions. **R**
  Add chlorine water to potassium bromide solution.
  Add chlorine water to potassium iodide solution.
  Add bromine water to potassium iodide solution.
- Demonstrate the formation of halides. Use small quantities in all of these reactions, which must be performed in an efficient fume cupboard.
  - Heat aluminium turnings in a current of chlorine or add hot aluminium powder to a gas jar of chlorine. Add hot aluminium powder to bromine vapour. Add one drop of water to an intimate mixture of aluminium powder and iodine.
  - Reaction of copper (use Dutch metal) and chlorine.

- Reaction of phosphorus and chlorine. Put a small piece of white phosphorus on a spoon into a gas jar of chlorine.
- Hydrogen and chlorine. Make a mixture of these gases by electrolysis and expose it to a bright light from a photographic flash or burning magnesium.

Metals

- Most metals react with dilute acids. Add a small piece of a metal to dilute hydrochloric acid and test for hydrogen. Use magnesium, aluminium, zinc, iron and copper.
- Differences in the physical properties of the transition elements, such as melting points and densities, can be established from data books.
- Transition metals react with oxygen to form insoluble oxides that are not alkaline. The reactions of these metals with acids will illustrate their lower reactivity.
- An observational exercise using compounds of chromium, manganese, cobalt, nickel and copper will emphasise that these metals have coloured compounds.
- Flame colours (p. 261).

Chapter 10 Metals

Typical reactions of metals

- Reaction of metals with dilute acids (p. 150). **R**
- Reaction of metals with air and oxygen (p. 151). **R**
- Reaction of metals with cold water or steam (p. 151). **R**

Competition reactions

- Reduction of metal oxides by other metals (Figure 10.6, p. 154). Demonstrate the Thermit reaction and the reactions between the metals magnesium, zinc, iron and copper and their oxides. This will establish an order of reactivity for these metals. Some of these reactions are very violent so the use of small quantities and a rehearsal before the class demonstration are essential. **R**
- Reduction of metal oxides by methane/hydrogen. It is safer to use methane, which is passed over heated copper oxide. **R**
- Metal/metal displacement reactions in aqueous solution. Clean pieces of a metal are added to the aqueous nitrate of another metal. The students look for evidence of displacement and so establish an order of reactivity. Suitable metals are magnesium, zinc, lead, copper and silver. **R**

Identifying metal ions

- Reactions of cations in aqueous solution with sodium hydroxide – Al^{3+}, Fe^{2+}, Fe^{3+}, Cu^{2+}, Ca^{2+}, Zn^{2}, Cr^{3+},. An excess of sodium hydroxide solution is added slowly to a small volume of the solution containing the cation. This is an exercise in observation, organisational and recording skills and in the ability to write chemical equations – word, molecular and ionic.
- Reactions of the same cations with aqueous ammonia. This is essentially a repeat of the exercise above but with less emphasis on equation writing.
- Identification of unknown cations using sodium hydroxide solution and aqueous ammonia.

Amphoteric hydroxides and oxides

- Hydroxides are classified as basic or amphoteric. The qualitative exercises above will provide a practical basis for the underpinning theory. A useful extension exercise is to provide unknown cations in solution and for the students to ascertain if their hydroxides are basic or amphoteric.
- Another informative practical involves adding dilute acid dropwise to an aluminate or zincate and observing the changes in reverse.

$$Al^{3+} \underset{H^+}{\overset{OH^-}{\rightleftharpoons}} Al(OH)_3 \underset{H^+}{\overset{OH^-}{\rightleftharpoons}} Al(OH)_4^-$$

- Oxides are classified as basic, acidic, neutral or amphoteric. An exercise with unknown oxides is conducted using the following tests:
  - Does the oxide dissolve in water? If so, measure the pH.
  - Do the insoluble oxides react with 4 mol dm^{-3} nitric acid?
  - Do the insoluble oxides react with 4 mol dm^{-3} sodium hydroxide?

Action of heat on metal compounds

- Action of heat on metal oxides.
- Action of heat on hydroxides. It is easier to use metal hydroxides that have a different colour from

their oxides, such as the hydroxides of iron, lead and copper.

- Action of heat on nitrates. For safety reasons this is best demonstrated as there are some serious hazards – molten potassium and sodium nitrates/nitrite, decrepitation with lead nitrate and the emission of the poisonous gas, nitrogen dioxide.
- Action of heat on carbonates – see p. 152.

Rusting of iron

- Experiment on rusting as described in the text (Figure 10.20, p. 162). **R**
- Set up a cell with two iron electrodes connected through a voltmeter and bubble oxygen onto one electrode. From the polarity of the cell, students can deduce the direction of the electron flow and analyse the cell reactions in terms of electron transfer. **R**
- Set up a zinc/iron cell to demonstrate sacrificial protection. Analyse the cell reactions as above. **R**

Chapter 11 Air and water

Air

- Find the percentage of oxygen in the air as described in the text (Figure 11.6, p. 174).

Ammonia gas

- Laboratory preparation of ammonia (Figure 11.16, p. 179).
- The fountain flask experiment (Figure 11.18, p. 180).
- Show that ammonia is a base. Neutralise acids by adding 2 mol dm^{-3} aqueous ammonia to dilute acids in the presence of universal indicator.
- Show that ammonia is a weak base. Measure the pH of 0.1 mol dm^{-3} aqueous ammonia and its conductivity. Compare the pH and conductivity with those of 0.1 mol dm^{-3} sodium hydroxide.
- Precipitation of metal hydroxides – see Chapter 10.
- Test for the ammonium ion (p. 262).

Reversible reactions

- Acid/base indicators.
- Action of heat on copper(II) sulfate-5-water.
- Add sodium hydroxide solution to a solution of a zinc, an aluminium or a lead salt. The hydroxide will precipitate and then dissolve in excess alkali to form an aluminate, zincate or plumbate. On the addition of dilute nitric acid the hydroxide will reappear, only to dissolve with the addition of more nitric acid.

$$\text{metallic cation} \underset{H^+}{\overset{OH^-}{\rightleftharpoons}} \text{metal hydroxide} \underset{H^+}{\overset{OH^-}{\rightleftharpoons}} \text{metallic anion}$$

Equilibria

- Ice/water and other changes of state.
- Saturated solutions in equilibrium with excess solute.
- Bismuth chloride and water. Prepare a solution of bismuth chloride in the minimum of hydrochloric acid and add to a large excess of water.

$$\underset{\text{white}}{BiCl_3 + H_2O \rightleftharpoons BiOCl(s)} + 2HCl$$

On the addition of hydrochloric acid, the white precipitate will become fainter and then disappear. The addition of sodium hydroxide solution will cause the precipitate to reappear. This illustrates different positions of the same equilibrium.

- Iodine will react with chlorine to form iodine monochloride, and with an excess of chlorine the following equilibrium will be established.

$$Cl_2 + \underset{\text{brown}}{ICl(l)} \rightleftharpoons \underset{\text{yellow}}{ICl_3(s)}$$

The removal of chlorine by blowing air through the U-tube or by inverting it results in the formation of the brown liquid, while the addition of chlorine results in the formation of the yellow solid.

- $Fe^{3+}(aq) + Ag(s) \rightleftharpoons Fe^{2+}(aq) + Ag^+(aq)$
- $FeCl_3 + 3KCNS \rightleftharpoons Fe(CNS)_3 + 3KCl$

Although this equation is not strictly correct, at this level it is preferable to the ionic equation involving $FeCNS^{2+}$.

- $N_2O_4 \rightleftharpoons 2NO_2$

Altering the temperature using ice-water or hot water will change the position of equilibrium and illustrate its dependence on the exothermic or endothermic nature of the reaction. Similarly, changing the pressure by compressing a sample of the equilibrium mixture in a gas syringe will also move the position of equilibrium.

- $NH_4Cl \rightleftharpoons NH_3 + HCl$

Heat ammonium chloride to show the influence of temperature on the position of equilibrium.

Water

- Tests to show the presence of water using cobalt chloride paper and anhydrous copper(II) sulfate.
- Tests to show that water is pure: it melts at 0 °C or boils at 100 °C.

Hardness in water

- Standard experiments with permanently and temporarily hard water. Measure the number of drops of soap solution needed to form a permanent lather in 10 cm^3 samples of both types of hard water. Boil both types and allow to cool. Then measure the number of drops of soap solution needed to produce a permanent lather in 10 cm^3 samples of the treated samples of hard water. The study can be extended to include the addition of sodium carbonate and treatment with ion exchange resins.
- Compare soap and soapless detergents. Add soap solution dropwise with shaking to samples of hard and soft water. Repeat using a soapless detergent.

Chapter 12 Sulfur

Sulfur and sulfur compounds

- Burn sulfur in air or oxygen. Show that an acidic gas is formed.
- Formation of metal sulfides. Demonstrate the reactions of sulfur with iron, zinc and magnesium – care is needed.
- Preparation of sulfur trioxide. Pass dry oxygen and sulfur dioxide over heated platinised mineral wool. Collect sulfur trioxide as a solid in a cooled receiver.
- Reaction of sulfur trioxide with water – care is needed. Test the solution to show that it is acidic (using both universal indicator and magnesium powder), and that it contains sulfate ions (by adding acidified barium chloride solution).

Dilute sulfuric acid

- Reaction with magnesium and zinc. **R**
- Reaction with bases – CuO and NaOH(aq).
- Reaction with carbonates – $NiCO_3$, $CoCO_3$ and Na_2CO_3(aq).
- Precipitation of insoluble sulfates of barium, lead and calcium.
- Preparation of sodium sulfate-10-water by titration. Use the same method as on p. 124, using dilute sulfuric acid and dilute aqueous sodium hydroxide.
- Preparation of sodium hydrogensulfate by titration. Use the results from the previous exercise but add double the titre of acid to the same volume of alkali.
- Comparison of the salts sodium sulfate and sodium hydrogensulfate. Compare their crystal shapes. Measure the pH of their aqueous solutions. Add magnesium powder to both salts in solution. Add excess dilute hydrochloric acid and barium chloride solution to both solutions.

Concentrated sulfuric acid

- Reactions with copper(II) sulfate-5-water and sugar.
- Oxidation of metals, such as copper.
- Make nitric acid and sulfuric acid.

Chapter 13 Inorganic carbon chemistry

Carbonates

- Heat a piece of calcium carbonate. Allow it to cool and carefully add water dropwise. Then add excess water and filter. Keep the filtrate, a solution of calcium hydroxide. Make carbon dioxide by adding an acid to calcium carbonate. Bubble this gas through the previously prepared calcium hydroxide solution until no further change is observed. Boil the resulting solution of calcium hydrogencarbonate.
- Action of heat, if any, on sodium, zinc and copper carbonates.
- Reaction of acids with a selection of carbonates.
- Formation of insoluble carbonates by precipitation. Add aqueous sodium carbonate to solutions of copper sulfate, calcium chloride and nickel sulfate.

Carbon dioxide

- Laboratory preparation of carbon dioxide and its properties (Figure 13.16, p. 214).
- Demonstrate that carbon and carbon-containing compounds form carbon dioxide on combustion. Burn a variety of materials and test for carbon dioxide with limewater. Suitable materials are a candle, wood shavings, charcoal, paraffin etc.

Chapter 14 Organic chemistry 1

- Demonstrate combustion of some alkanes. Use a Bunsen burner to show complete and incomplete combustion of methane. Burn a range of alkanes to show the variation in ease of ignition – pentane and hexane are highly flammable but liquid paraffin and paraffin wax need pre-heating and/or a wick.
- Reaction of alkenes with bromine.
- Cracking of alkanes as described in the text (Figure 14.11, p. 224).
- Test for unsaturation using bromine in a suitable solvent (Figure 14.13, p. 225).
- Demonstration of thermosoftening and thermosetting plastics. Heat small samples of poly(ethene) and melamine.

Chapter 15 Organic chemistry 2

Ethanol

- Combustion of ethanol in a small spirit lamp to demonstrate that it burns with a small blue flame.
- Add ethanol to water to show that the two liquids are miscible.
- Show examples of ethanol as a solvent.
- Oxidation of ethanol by the air. Use a water pump to bubble air through dilute aqueous ethanol. **R**
- Add ethanol dropwise to a warm mixture of potassium dichromate(VI) and dilute sulfuric acid. **R**
- Formation of esters using ethanol and higher alcohols. Mix 1 cm^3 of ethanol and 1 cm^3 of glacial ethanoic acid. Add 2 drops of concentrated sulfuric acid and warm the mixture for a few minutes. Then pour the mixture into a beaker of cold water. The distinctive smell of an ester can be detected. Repeat the procedure using higher alcohols – propanol, butanol and pentanol.
- Making ethanol by fermentation (Figure 15.8, p. 236).
- Fractional distillation of a mixture of water and ethanol, using the apparatus shown in Figure 2.22 (p. 21). Note that this procedure may be illegal in some countries.

Ethanoic acid

- Show that ethanoic acid is an acid.
  - Reaction with metals – magnesium and zinc powders.
  - Reaction with carbonates – aqueous sodium carbonate and copper(II) carbonate.
  - Reaction with bases – sodium hydroxide solution, copper(II) oxide and magnesium oxide.
  - Reaction with an indicator – litmus.
- Show that ethanoic acid is a weak acid – a comparative study with hydrochloric acid.
  - Measure the pH of 0.1 mol dm^{-3} solutions of both acids.
  - Measure the conductivity of 0.1 mol dm^{-3} solutions of both acids.
  - Measure the time taken for a short length of magnesium ribbon to react with 1.0 mol dm^{-3} solutions of both acids.
- Formation of esters (p. 237).

Soaps and detergents

- Making soap from vegetable oil and alkali. Boil a vegetable oil or animal fat with aqueous sodium hydroxide. Cool and add salt to precipitate the soap.
- Preparation of a soapless detergent from a vegetable oil.

Amino acids

- Identification of amino acids by chromatography. Use a 4 : 1 : 2 mixture of butanol, ethanoic acid and water as a solvent. Dry the chromatogram and develop with a ninhydrin spray.
- Biuret test for proteins. Warm the sample in water, add concentrated sodium hydroxide solution and a drop of very dilute copper(II) sulfate solution.

Plastics and polymers

- The 'nylon rope trick'. The diamine, 1,6-diaminohexane, is dissolved in water to which some sodium carbonate has been added. A solution of a diacid chloride, adipoyl chloride, in cyclohexane is added and a nylon thread can be pulled from the interface between the two phases.
- Making a solid foam. The ingredients for making a polyurethane foam can be purchased as a demonstration kit.

Starch and glucose

- Tests for starch and glucose. Test for starch using dilute aqueous iodine and for glucose with Fehling's solution or Benedict's reagent.
- Hydrolysis of starch (p. 242). Use the tests above to show that glucose is formed and starch is used up when aqueous starch is boiled with dilute hydrochloric acid.

Notes on qualitative analysis

The branch of chemistry that deals with the identification of elements or grouping of elements present in a sample is called **qualitative chemical analysis**, or **qualitative analysis** for short. It does not deal with anything to do with quantities.

The techniques employed in qualitative analysis vary in their complexity, depending on the nature of the sample under investigation. In some cases it is only necessary to confirm the presence of certain elements or groups for which specific chemical tests, or 'spot' tests, applicable directly to the sample, may be available. More often, the sample is a complex mixture, and a systematic analysis must be made in order that all the component parts may be identified. Often, the first simple stages of qualitative analysis require no apparatus at all. Things like colour and smell can be observed without any need for apparatus.

The following summary collects together information from throughout the book which would allow you to carry out qualitative analysis.

Appearance or smell

A preliminary examination of the substance will give you a start. The appearance or smell of a substance can often indicate what it might contain (see Table 16.1).

Table 16.1 Deductions that can be made from a substance's appearance or smell.

Observation on substance	Indication
Black powder	Carbon, or contains O^{2-} ions (as in CuO), or S^{2-} ions (as in CuS)
Pale green crystals	Contains Fe^{2+} ions (as in iron(II) salts)
Dark green crystals	Contains Ni^{2+} ions (as in nickel(II) salts)
Blue or blue-green crystals	Contains Cu^{2+} ions (as in copper(II) salts)
Yellow-brown crystals	Contains Fe^{3+} ions (as in iron(III) salts)
Smell of ammonia	Contains NH_4^+ ions (as in ammonium salts)

Flame colours

If a clean nichrome wire is dipped into a metal compound and then held in the hot part of a Bunsen flame, the flame can become coloured (Figure 16.3). Certain metal ions may be detected in their compounds by observing their flame colours (Table 16.2).

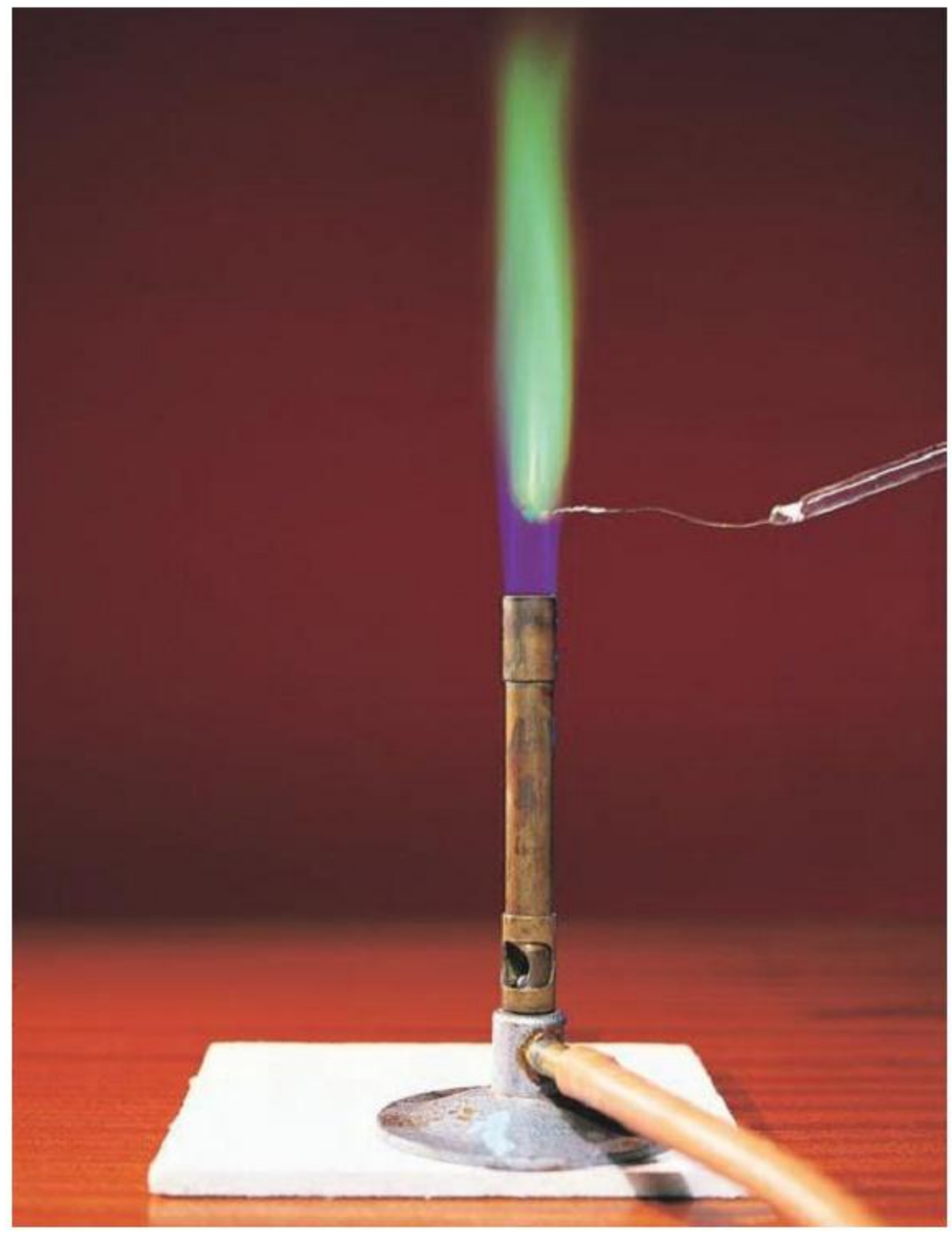


Figure 16.3 The green colour is characteristic of copper.

Table 16.2 Characteristic flame colours of some metal ions.

	Metal	Flame colour
Group I	Lithium	Crimson
	Sodium	Golden yellow
	Potassium	Lilac
	Rubidium	Red
	Caesium	Blue
Group II	Calcium	Brick red
	Strontium	Crimson
	Barium	Apple green
Others	Lead	Blue-white
	Copper (as Cu(II))	Green

A flame colour is obtained as a result of the electrons in the particular ions being excited when they absorb energy from the flame which is then emitted as visible light. The different electronic structures of the different ions, therefore, give rise to the different colours.

Tests for aqueous cations

Effect of adding dilute sodium hydroxide solution

Aqueous sodium hydroxide can be used to identify salts of Al^{3+}, Ca^{2+}, Cr^{3+}, Cu^{2+}, Fe^{2+}, Fe^{3+}, Pb^{2+} and Zn^{2+} when present in aqueous solutions. All metal cations form insoluble hydroxides when sodium hydroxide solution is added to them. The colour of the precipitate and its behaviour in excess sodium hydroxide solution will help identify the metal present (Table 16.3).

Table 16.3 Effect of adding sodium hydroxide solution to solutions containing various metal ions.

Added dropwise	To excess	Likely cation
White precipitate	Precipitate dissolves	Al^{3+}, Zn^{2+} or Pb^{2+}
White precipitate	Precipitate does not dissolve	Ca^{2+}
Blue precipitate	Precipitate does not dissolve	Cu^{2+}
Green precipitate	Precipitate does not dissolve	Fe^{2+}
Brown precipitate	Precipitate does not dissolve	Fe^{3+}
Light green precipitate	Precipitate dissolves	Cr^{3+}

(In the case of ammonium salts, ammonia gas is produced on warming. The ammonium cation does not form an insoluble hydroxide. However, it forms ammonia and water upon heating.)

Effect of adding dilute ammonia solution

Ammonia gas dissolved in water is usually known as aqueous ammonia. The solution is only weakly alkaline, which results in a relatively low concentration of hydroxide ions. Aqueous ammonia can be used to identify salts of Al^{3+}, Ca^{2+},Cr^{3+}, Cu^{2+}, Fe^{2+}, Fe^{3+}, Pb^{2+} and Zn^{2+} ions. The colour of the precipitate or solution formed identifies the metal present (Table 16.4).

Table 16.4 Effect of adding aqueous ammonia to solutions containing various metal ions.

Added dropwise	To excess	Cation present
Gelatinous blue precipitate	Precipitate dissolves to give a deep blue clear solution	Cu^{2+}
Dirty green precipitate	Precipitate does not dissolve	Fe^{2+}
Rust brown precipitate	Precipitate does not dissolve	Fe^{3+}
White precipitate	Precipitate does not dissolve	Al^{3+}
White precipitate	Precipitate dissolves	Zn^{2+}
No precipitate	No precipitate	Ca^{2+}
Light green precipitate	Precipitate dissolves to some extent	Cr^{3+}

Tests for aqueous anions

Table 16.5 shows a variety of tests for aqueous anions.

Table 16.5 Tests for aqueous anions.

Anion	Test	Test result
Carbonate (CO_3^{2-})	Add dilute acid	Effervescence, carbon dioxide produced
Chloride (Cl^-) [in solution]	Acidify with dilute nitric acid, then add aqueous silver nitrate	White ppt.
Bromide (Br^-) [in solution]	Acidify with dilute nitric acid, then add aqueous silver nitrate	Cream ppt.
Iodide (I^-) [in solution]	Acidify with dilute nitric acid, then add aqueous silver nitrate	Yellow ppt.
Nitrate (NO_3^-) [in solution]	Add aqueous sodium hydroxide, then aluminium foil; warm carefully	Ammonia produced
Sulfate (SO_4^{2-}) [in solution]	Acidify, then add aqueous barium nitrate or barium chloride	White ppt.
Sulfite (SO_3^{2-})	Add dilute hydrochloric acid, warm gently and test for the presence of sulfur dioxide	Sulfur dioxide produced will turn acidified aqueous potassium manganate(VII) from purple to colourless

Tests for gases

Table 16.6 shows the common gases which may be produced during qualitative analysis and tests which can be used to identify them. These tests are used in conjunction with the tests shown above.

Table 16.6 Tests for gases.

Gas	Colour (odour)	Effect of moist indicator paper	Test
Hydrogen (H_2)	Colourless (odourless)	No effect – neutral	'Pops' in the presence of a lighted splint
Oxygen (O_2)	Colourless (odourless)	No effect – neutral	Relights a glowing splint
Carbon dioxide (CO_2)	Colourless (odourless)	Pink – weakly acidic	Turns limewater a cloudy white
Ammonia (NH_3)	Colourless (very pungent smell)	Blue – alkaline	Turns moist indicator paper blue – it is the only alkaline gas
Sulfur dioxide (SO_2)	Colourless (very choking smell)	Red – acidic	• Turns acidified potassium dichromate(VI) from orange to green • Turns acidified potassium manganate(VII) from purple to colourless
Chlorine (Cl_2)	Yellow-green (very choking smell)	Bleaches moist indicator paper after it initially turns pale pink	Bleaches moist indicator paper
Nitrogen(IV) oxide (NO_2)	Brown (very choking smell)	Pink – weakly acidic	• Brown colour • Turns moist indicator paper pink
Water (H_2O)	Colourless (odourless)	No effect – neutral	• Turns blue cobalt chloride paper pink • Turns anhydrous copper(II) sulfate from white to blue

Questions to help your understanding

1 For each of the following pairs of substances, describe a chemical test you would carry out to distinguish between them.
   a potassium sulfate and potassium sulfite
   b ammonium chloride and aluminium chloride
   c zinc nitrate and calcium nitrate
   d sodium chloride and sodium iodide
   e iron(II) sulfate and copper(II) sulfate

2 Sodium carbonate hydrate contains water of crystallisation. When it is heated strongly it gives off the water of crystallisation, which can be collected.
   a The substance left behind is anhydrous sodium sulfate. Describe a chemical test to show that this substance contains sodium (cation) and carbonate (anion).
   b Describe two chemical tests to show that the colourless liquid produced and collected is indeed water.
   c Describe another test to show that the colourless liquid given off in this experiment is water.

Revision and exam-style questions

Do not write in this book. Where necessary copy drawings, tables or sentences.

Alternative to practical paper

1 A student investigated the reaction of aqueous sodium hydroxide with two different acids, acid **C** and acid **D**.

Two experiments were carried out.

Experiment 1

By using a measuring cylinder, 20 cm^3 of aqueous sodium hydroxide was poured into a conical flask and the initial temperature of the solution was measured.

A burette was filled with acid **C** up to the 0.0 cm^3 mark.

5 cm^3 of acid **C** was added to the sodium hydroxide in the flask. The temperature of the mixture was measured.

Further 5 cm^3 portions of acid **C** were added to the mixture in the flask, stirring with the thermometer until a total volume of 30 cm^3 of acid **C** had been added. The temperatures after each 5 cm^3 portion had been added were measured.

a Use the thermometer diagrams to record the temperatures in the table of results.

Table of results

Volume of acid C added/cm^3	Thermometer diagrams	Temperature/°C
0	30 25 20	
5	40 35 30	
10	40 35 30	
15	40 35 30	
20	35 30 25	
25	35 30 25	
30	35 30 25	

[2]

Experiment 2

The burette was emptied and rinsed with water. Experiment 1 was repeated using acid **D**.

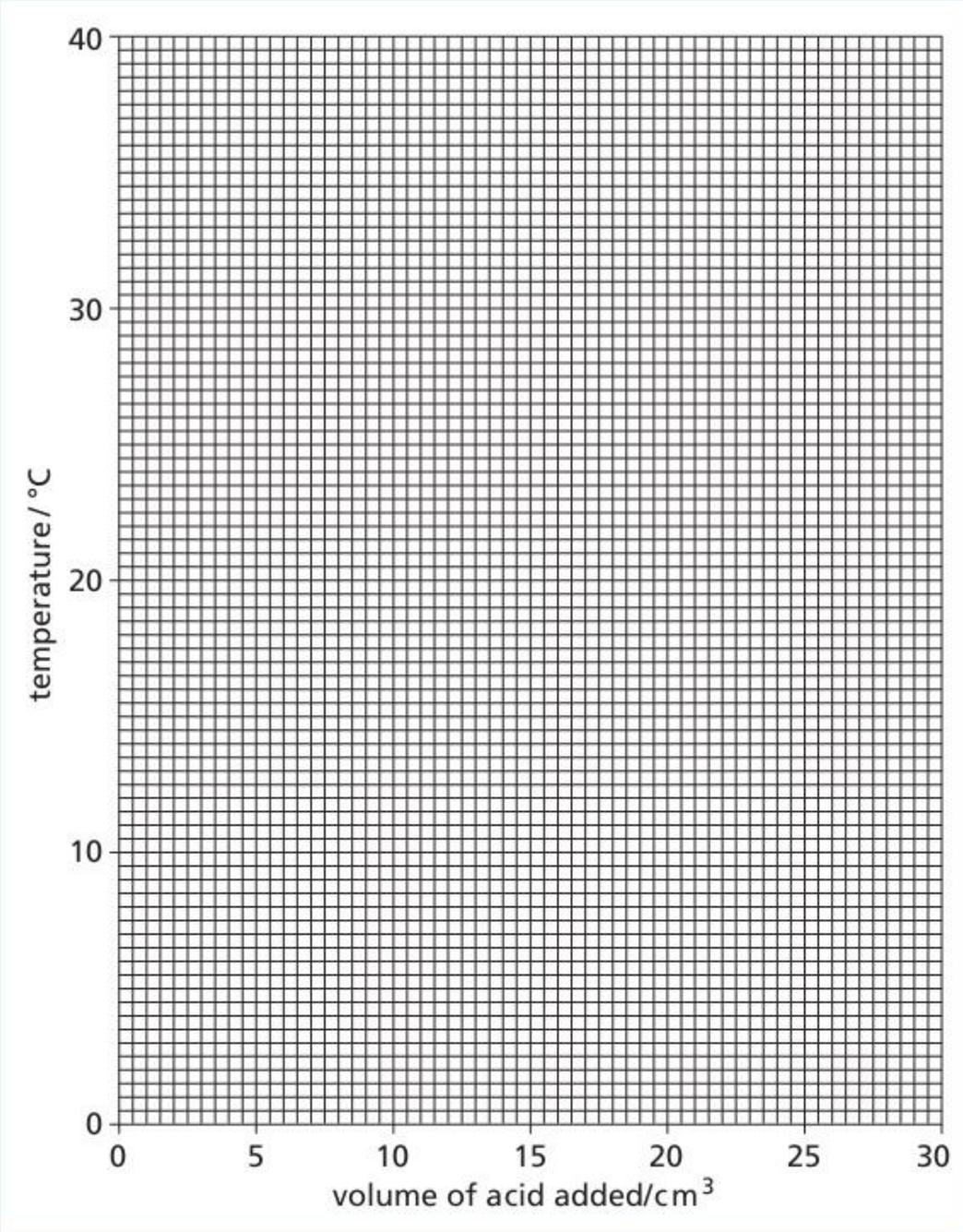
b Use the thermometer diagrams to record the temperatures in the table of results.

Table of results

Volume of acid D added/cm^3	Thermometer diagrams	Temperature/°C
0	30 25 20	
5	30 25 20	
10	30 25 20	
15	30 25 20	
20	30 25 20	
25	30 25 20	
30	30 25 20	

[2]

c Plot the results for Experiments 1 and 2 on the grid and draw two smooth line graphs. Clearly label your graphs.



[6]

d From your graph, deduce the temperature of the mixture when 3 cm^3 of acid **C** reacted with sodium hydroxide in Experiment 1.
Show clearly **on the graph** how you worked out your answer. [2]

e **(i)** Which experiment produced the larger temperature change? [1]
(ii) Suggest why the temperature change is greater in this experiment. [2]

f Why was the burette rinsed with water in Experiment 2? [1]

g Predict the temperature of the reaction mixture in Experiment 2 after 1 hour. Explain your answer. [2]

[Total: 18]

(Cambridge IGCSE Chemistry 0620 Paper 61 Q4 June 2010)

2 Solid **E** was analysed. **E** was an aluminium salt. The tests on the solid and some of the observations are in the following table. Complete the observations in the table.

Tests	Observations
Tests on solid E	
a Appearance of solid **E**	white crystalline solid
b A little of solid **E** was heated in a test-tube.	colourless drops of liquid formed at the top of the tube
c A little of solid E was dissolved in distilled water. The solution was divided into four test-tubes and the following tests were carried out.	
(i) To the first test-tube of solution, drops of aqueous sodium hydroxide were added. Excess sodium hydroxide was then added to the test-tube.	.... [3]
(ii) Test (i) was repeated using aqueous ammonia solution instead of aqueous sodium hydroxide.	.... [2]
(iii) To the third test-tube of solution, dilute hydrochloric acid was added, followed by barium chloride solution.	no reaction
(iv) To the fourth test-tube of solution, aqueous sodium hydroxide and aluminium powder were added. The mixture was heated.	effervescent pungent gas given off turned damp litmus paper blue

d What does test **(b)** tell you about solid **E**? [1]

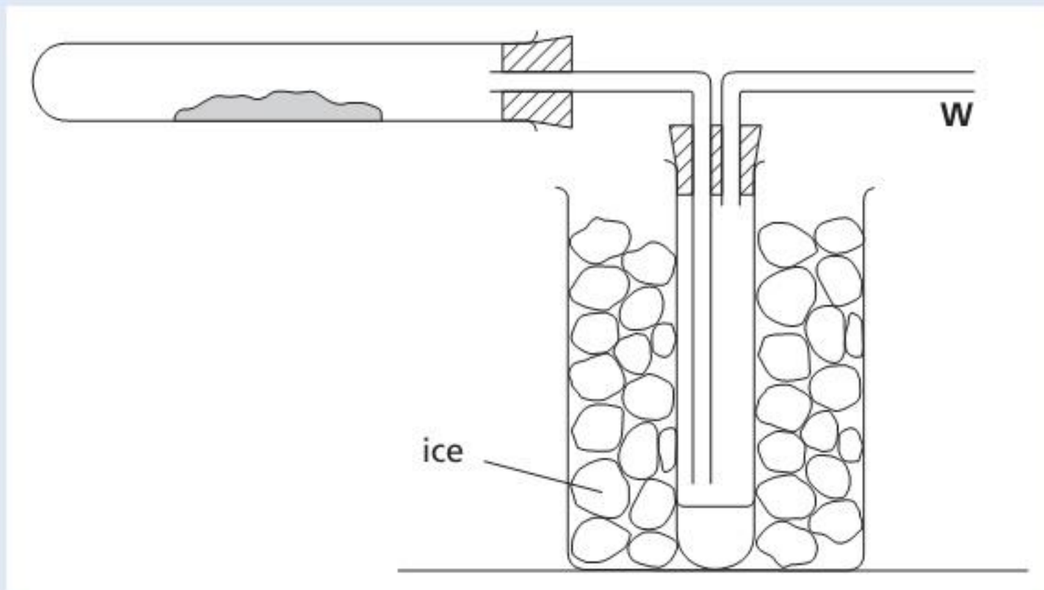
e Identify the gas given off in test **(c)(iv)**. [1]

f What conclusions can you draw about solid **E**? [2]

[Total: 9]

(Cambridge IGCSE Chemistry 0620 Paper 61 Q5 June 2010)

3 Hydrated cobalt chloride crystals, $CoCl_2.6H_2O$, were heated in the apparatus shown below.



a Indicate on the diagram, using an arrow, where heat is applied. [1]
b The crystals change colour from _____ to _____ . [1]
c What is the purpose of the ice? [1]
d Why is the tube open at point **W**? [1]

[Total: 4]

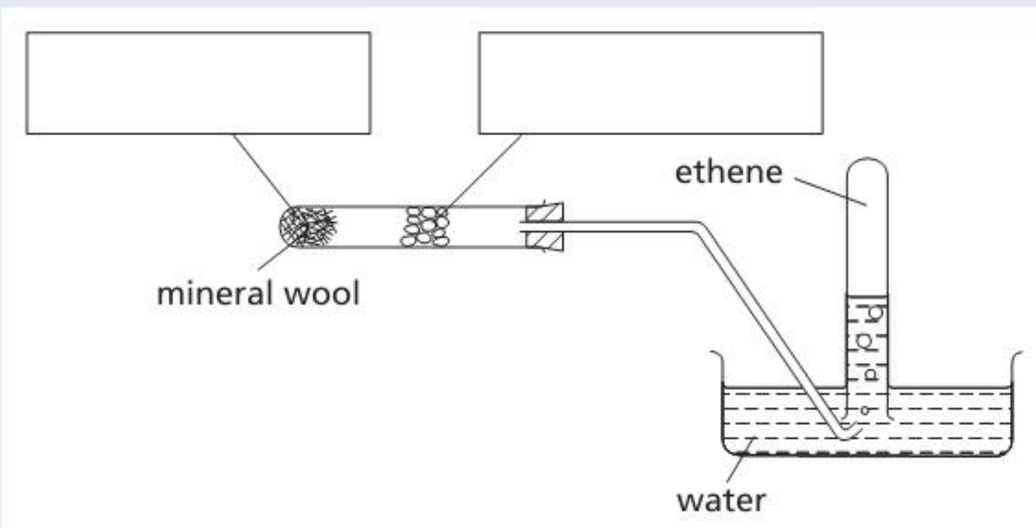
(Cambridge IGCSE Chemistry 0620 Paper 61 Q6 June 2010)

4 Malachite is a naturally occurring form of copper carbonate. Outline how a sample of copper metal could be obtained from large lumps of malachite in the laboratory.
Copper is one of the least reactive metals. Your answer should include any chemicals used and conditions. [6]

[Total: 6]

(Cambridge IGCSE Chemistry 0620 Paper 61 Q7 June 2010)

5 Ethene can be prepared by passing ethanol vapour over hot aluminium oxide.



a Complete the boxes to show the chemicals used. [1]
b Show on the diagram with two arrows where the heat is applied. [2]
c Why must the delivery tube be removed from the water before the heating is stopped? [2]

[Total: 5]

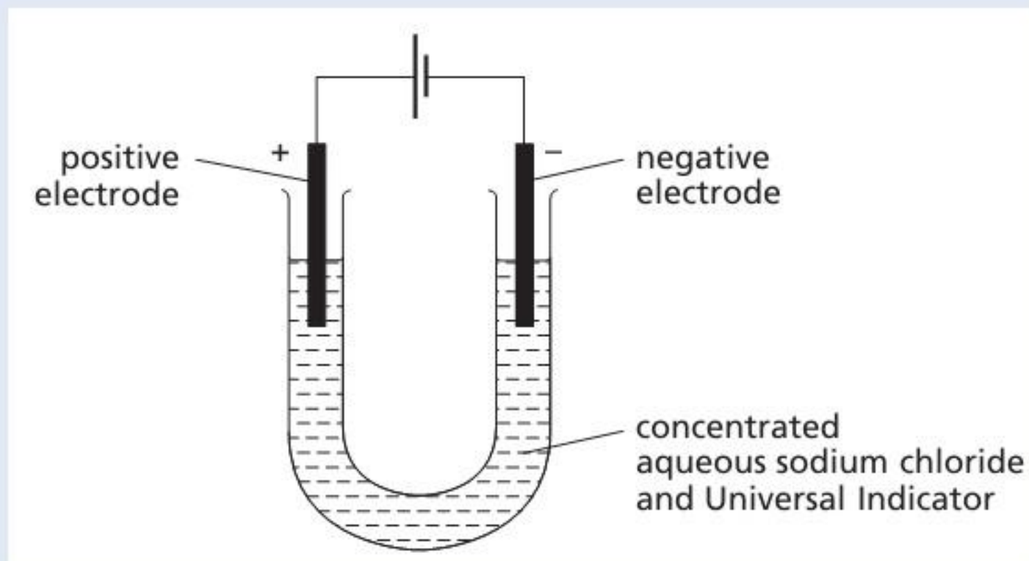
(Cambridge IGCSE Chemistry 0620 Paper 61 Q1 November 2010)

6 The following instructions were used to prepare magnesium sulfate crystals, $MgSO_4.7H_2O$.
Step 1 Measure 50 cm^3 of dilute sulfuric acid into a beaker and warm the solution.
Step 2 Using a spatula, add some magnesium oxide and stir the mixture. Continue adding the magnesium oxide until excess is present.
Step 3 Separate the excess magnesium oxide from the solution of magnesium sulfate.
Step 4 Heat the solution until crystals form. Obtain the crystals and dry them.
a Why is the sulfuric acid warmed? [1]
b How would you know when excess magnesium oxide is present in **Step 2**? [1]
c What method is used in **Step 3**? [1]
d Why must care be taken when drying the crystals in **Step 4**? [1]
e Explain how the method would differ if magnesium carbonate was used instead of magnesium oxide. [2]

[Total: 6]

(Cambridge IGCSE Chemistry 0620 Paper 61 Q2 November 2010)

7 A concentrated solution of sodium chloride was electrolysed using the apparatus below.



One observation noted was that the Universal Indicator turned purple at the negative electrode.
a What observation would be made at **both** electrodes? [1]
b Why did the indicator turn purple at the negative electrode? [1]
c **(i)** Name the product formed at the positive electrode. [1]
(ii) Suggest the effect of this product on the Universal Indicator. [1]

[Total: 4]

(Cambridge IGCSE Chemistry 0620 Paper 61 Q6 November 2010)

8 E numbers identify chemicals which are added to foods.

a E210 is benzoic acid. How could you show that a solution of benzoic acid is a weak acid?
test _____
result _____ [2]

b E211 is sodium benzoate. Name a suitable substance that would react with a solution of benzoic acid to form sodium benzoate. [1]

c E110 is Sunset Yellow. Outline a method you could use to show the presence of E110 in a food colouring. A space has been left if you want to draw a diagram to help answer the question. [4]

[Total: 7]

(Cambridge IGCSE Chemistry 0620 Paper 61 Q7 November 2010)

9 Describe a chemical test to distinguish between each of the following pairs of substances.
An example is given.
Example: hydrogen and carbon dioxide
test *lighted splint*
result with hydrogen *gives a pop*
result with carbon dioxide *splint is extinguished*

a zinc carbonate and zinc chloride
test _____
result with zinc carbonate _____
result with zinc chloride _____ [2]

b ammonia and chlorine
test _____
result with ammonia _____
result with chlorine _____ [3]

c aqueous iron(II) sulfate and aqueous iron(III) sulfate
test _____
result with aqueous iron(II) sulfate _____
result with aqueous iron(III) sulfate _____ [3]

[Total: 8]

(Cambridge IGCSE Chemistry 0620 Paper 06 Q3 June 2009)

10 Two solids, **S** and **V**, were analysed. **S** was copper(II) oxide.
The tests on the solids, and some of the observations, are in the following table.
Complete the observations in the table. Do not write any conclusions in the table.

Test	Observation
Tests on solid S	
a Appearance of solid **S**	black solid
b Hydrogen peroxide was added to solid **S** in a test-tube.	slow effervescence
A glowing splint was inserted into the tube.	splint relit
c Dilute sulfuric acid was added to solid S in a test-tube. The mixture was heated to boiling point.	blue solution formed
The solution was divided into three equal portions into test-tubes.	
(i) To the first portion of the solution, excess sodium hydroxide was added.	..... [1]
(ii) To the second portion of the solution, about 1 cm of aqueous ammonia solution was added.	..... [2]
Excess ammonia solution was then added.	.... [2]
(iii) To the third portion of the solution, dilute hydrochloric acid was added followed by barium chloride solution.	..... [2]
Tests on solid V	
d Appearance of solid **V**	black solid
e Hydrogen peroxide was added to solid **V** in a test-tube.	rapid effervescence
A glowing splint was inserted into the tube.	splint relit

f (i) Compare the reactivity of solid **S** and solid **V** with hydrogen peroxide. [1]
(ii) Identify the gas given off in test **(e)**. [1]

g What conclusions can you draw about solid **V**? [2]

[Total: 11]

(Cambridge IGCSE Chemistry 0620 Paper 06 Q5 June 2009)

11 Acid base indicators
Indicators are used to identify acids and bases.
Indicators can be obtained from berries and other fruits.

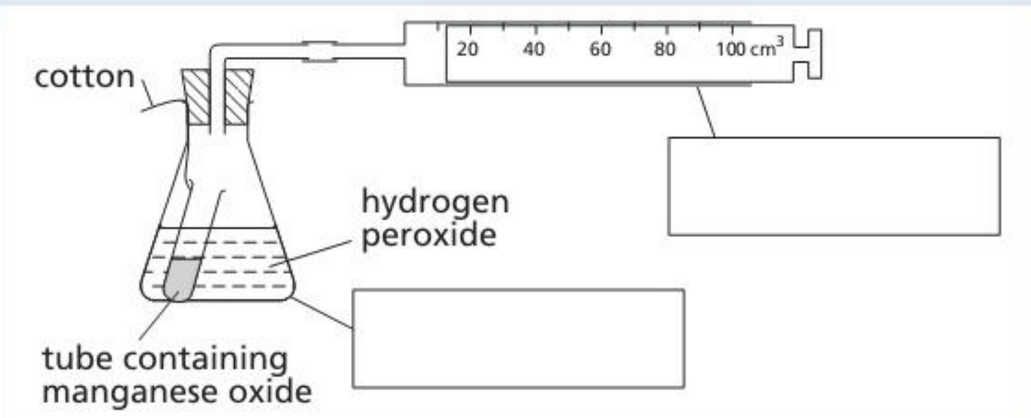
a Plan an experiment to obtain an aqueous solution of an indicator from some berries. [3]

b Plan an experiment to use the indicator solution to show that it is an effective indicator. [3]

[Total: 6]

(Cambridge IGCSE Chemistry 0620 Paper 06 Q6 June 2009)

12 The apparatus below was used to make oxygen. The tube of manganese oxide was added to the hydrogen peroxide solution by releasing the cotton.



a Complete the boxes to identify the pieces of apparatus. [2]

b Why was the tube of manganese oxide suspended in the flask? [1]

c Give a test for oxygen.
test _____
result _____ [2]

[Total: 5]

(Cambridge IGCSE Chemistry 0620 Paper 06 Q1 November 2009)

13 Three unlabelled bottles of chemicals each contained one of the following liquids:

- sodium nitrate dissolved in water;
- pure water;
- hexene.

a Give a test by which you could identify sodium nitrate solution.
test _____
result _____ [2]

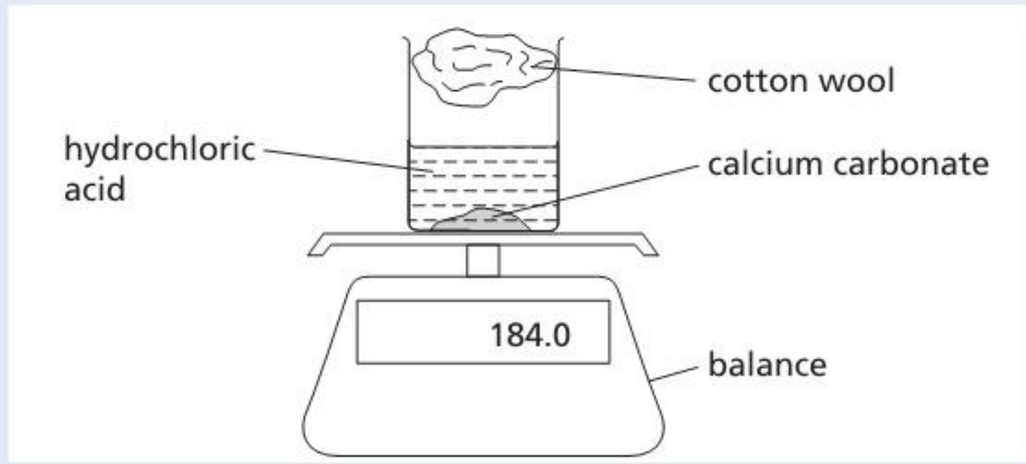
b Give a test by which you could identify pure water.
test _____
result _____ [2]

c Give a test by which you could identify hexene.
test _____
result _____ [2]

[Total: 6]

(Cambridge IGCSE Chemistry 0620 Paper 06 Q3 November 2009)

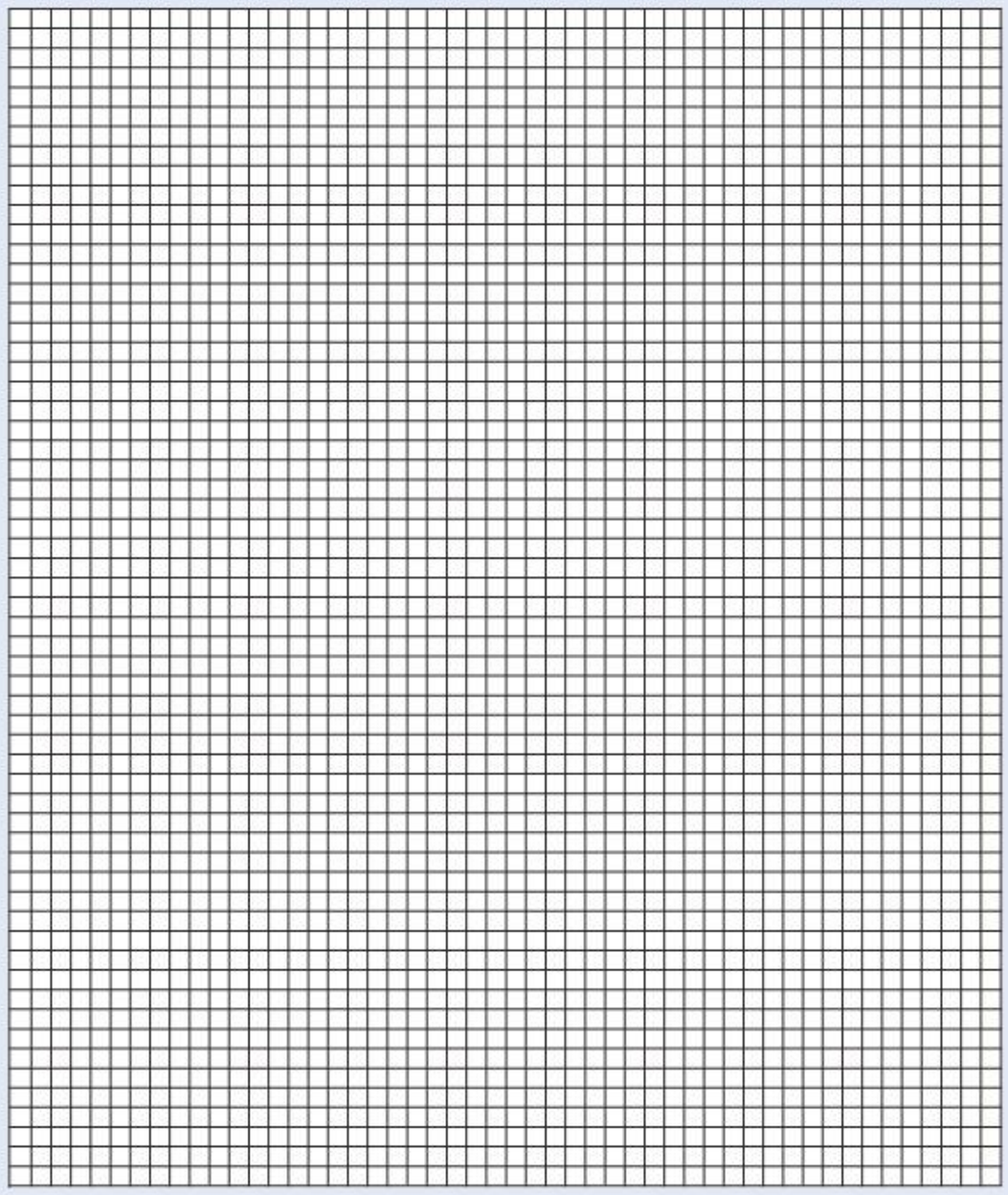
14 Dilute hydrochloric acid was added to excess calcium carbonate in a beaker as shown.



The beaker was placed on a balance and the mass of the beaker and contents recorded every minute.
The results are shown in the table.

Mass of beaker and contents/g	184.0	178.0	175.6	174.6	174.0	174.0
Time/min	0	1	2	3	4	5

a Plot the results on the grid and draw a smooth line graph.



[5]

b Use your graph to determine the mass of the beaker and contents after 30 seconds. Show clearly on your graph how you worked out your answer. [2]

c Why does the mass of the beaker and contents decrease? [1]
d Suggest the purpose of the cotton wool. [1]
e After how long did the reaction finish? [1]
f A second experiment was carried out using hydrochloric acid at a lower temperature. On the grid sketch a curve to show the expected results for this experiment. Label this curve **C**. [2]

[Total: 12]
(Cambridge IGCSE Chemistry 0620 Paper 06 Q6 November 2009)

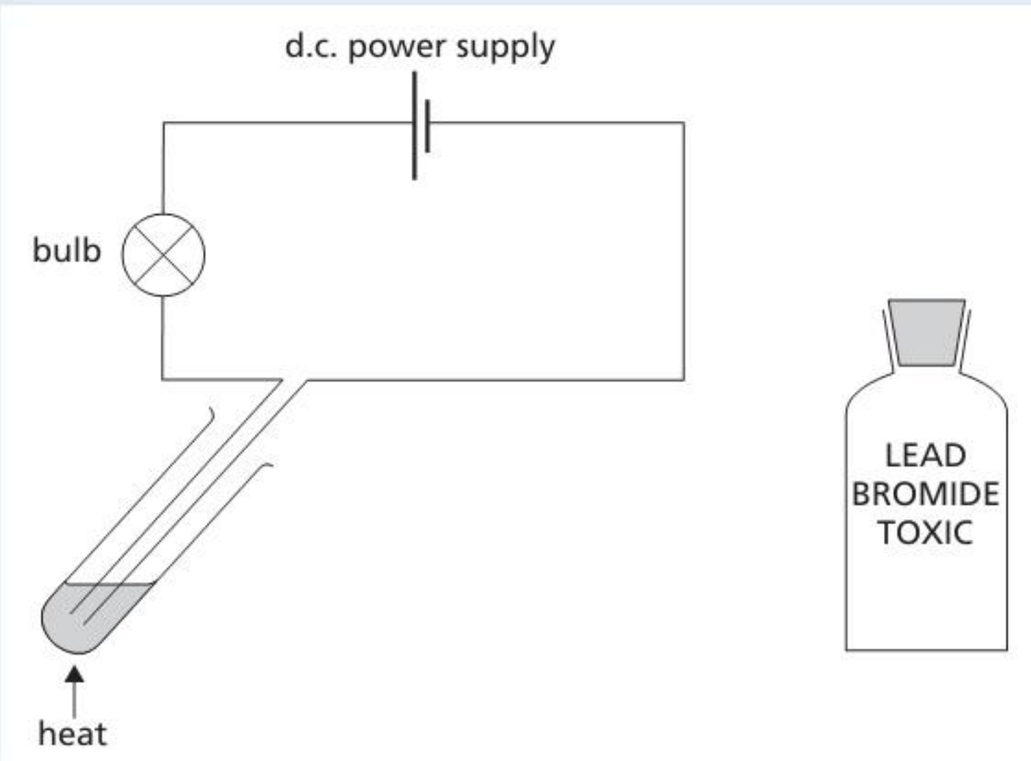
15 Leaves from trees contain a mixture of coloured pigments which are not soluble in water. A student was given these two instructions to investigate the pigments in the leaves.

1 Crush some leaves to extract the coloured pigments.
2 Use the liquid extract to find the number of coloured pigments in the leaves.

a What would the student need in order to effectively carry out instruction 1? [3]
b Describe an experiment to carry out instruction 2. A space has been left below if you want to draw a diagram to help answer the question. [4]

[Total: 7]
(Cambridge IGCSE Chemistry 0620 Paper 06 Q7 November 2009)

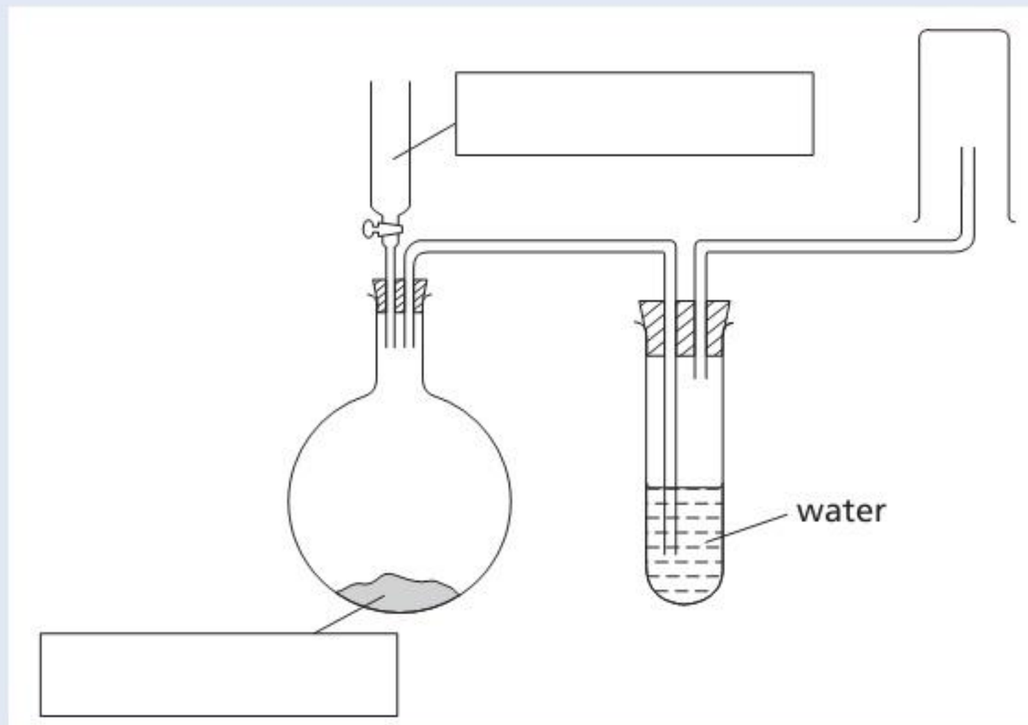
16 The diagram shows an experiment to pass electricity through lead bromide. Electricity has no effect on solid lead bromide.



a (i) Clearly label the electrodes on the diagram. [1]
(ii) Suggest a suitable material to make the electrodes. [1]
b Give two observations expected when the lead bromide is heated to melting point. [2]
c State two different safety precautions when carrying out this experiment. [2]

[Total: 6]
(Cambridge IGCSE Chemistry 0620 Paper 06 Q2 June 2008)

17 Sulfur dioxide gas is denser than air and soluble in water. A sample of sulfur dioxide can be prepared by adding dilute hydrochloric acid to sodium sulfite and warming the mixture. Study the diagram of the apparatus used.

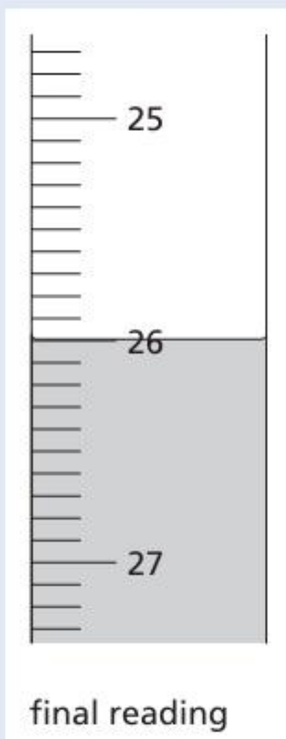


a Fill in the boxes to show the chemicals used. [2]
b Show by using an arrow, on the diagram, where heat is applied. [1]
c Identify and explain two mistakes in the diagram. [2]

[Total: 5]
(Cambridge IGCSE Chemistry 0620 Paper 06 Q3 June 2008)

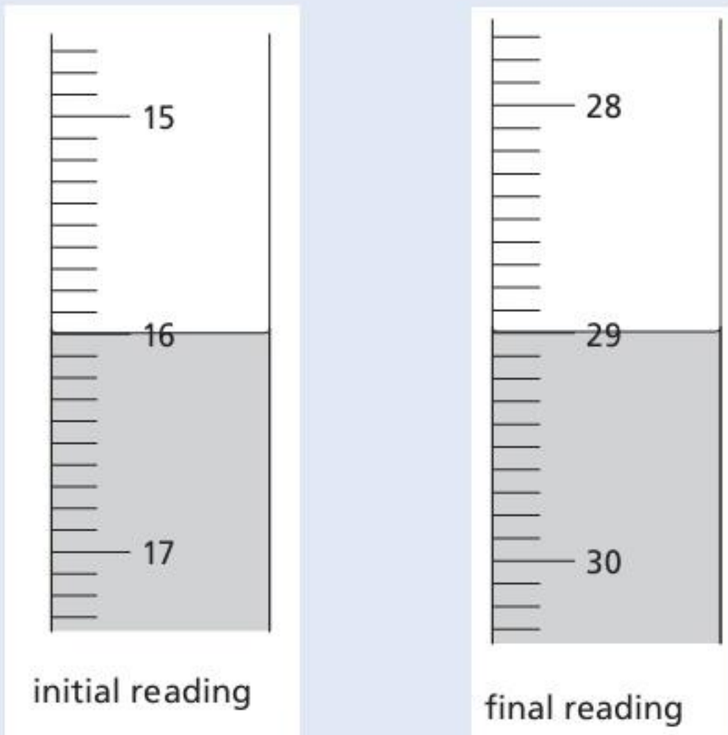
18 A student investigated the reaction between potassium manganate(VII) and a metallic salt solution.
Two experiments were carried out.
Experiment 1
a About 1 cm^3 of aqueous sodium hydroxide was added to a little of the salt solution **A** and the observation noted.
observation *green precipitate formed*

b A burette was filled with potassium manganate(VII) solution up to the 0.0 cm^3 mark. By using a measuring cylinder, 25 cm^3 of solution **A** of the salt was placed into a conical flask. The flask was shaken to mix the contents. The potassium manganate(VII) solution was added to the flask, and shaken to mix thoroughly. Addition of potassium manganate(VII) solution was continued until there was a pale pink colour in the contents of the flask.
Use the burette diagram to record the volume in the table and complete the column.



Experiment 2

c Experiment **1b** was repeated using a different solution **B** of the salt, instead of solution **A**. Use the burette diagrams to record the volumes in the table and complete the table.



d About 1 cm^3 of aqueous sodium hydroxide was added to a little of the solution in the flask and the observation noted.
observation *red-brown precipitate*

Table of results
Burette readings/cm^3

	Experiment 1	**Experiment 2**
Final reading		
Initial reading		
Difference		

[4]

e (i) In which experiment was the greatest volume of potassium manganate(VII) solution used? [1]
(ii) Compare the volumes of potassium manganate(VII) solution used in Experiments 1 and 2. [2]
(iii) Suggest an explanation for the difference in the volumes. [2]

f Predict the volume of potassium manganate(VII) solution which would be needed to react completely with 50 cm^3 of solution **B**. [2]

g Explain one change that could be made to the experimental method to obtain more accurate results.
change _____
explanation_____ [2]

h What conclusion can you draw about the salt solution from
(i) experiment **1a**, [1]
(ii) experiment **2d**? [1]

[Total: 15]

(Cambridge IGCSE Chemistry 0620 Paper 06 Q4 June 2008)

19 This label is from a container of 'Bite Relief' solution.

BITE RELIEF
FOR FAST RELIEF FROM INSECT BITES AND STINGS

Active ingredient: Ammonia
Also contains water and alcohol

DIRECTIONS FOR USE: Use cotton wool to dab the solution on the affected area of the skin

a Give a chemical test to show the presence of ammonia in Bite Relief solution.
test _____
result _____ [2]

b What practical method could be used to separate the mixture of alcohol (bp 78 °C) and water (bp 100 °C)? [2]

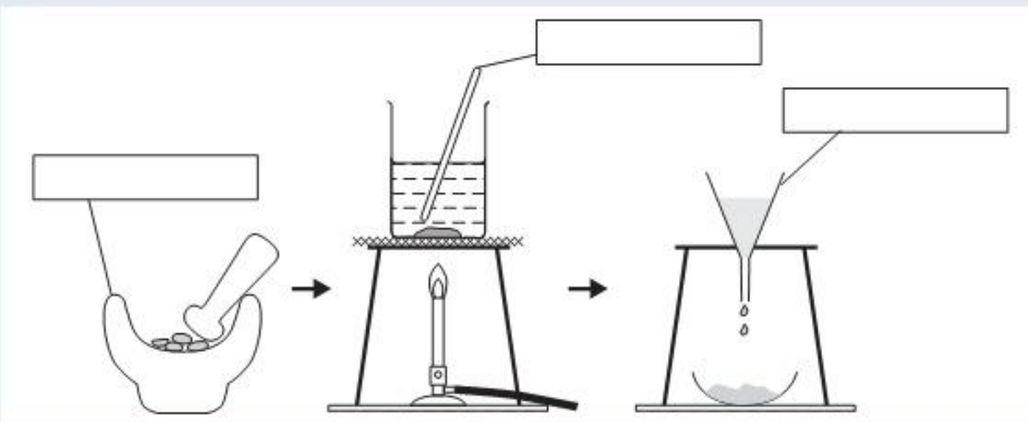
c Give a chemical test to show the presence of water.
test _____
result _____ [2]

d What would be the effect of touching the alcohol with a lighted splint? [1]

[Total: 7]

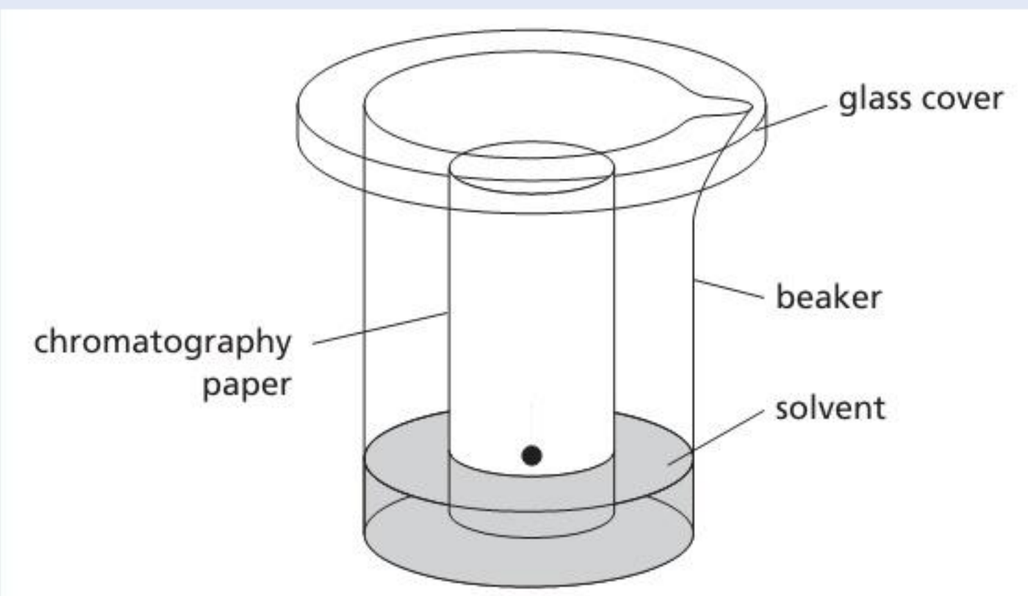
(Cambridge IGCSE Chemistry 0620 Paper 06 Q7 June 2008)

20 The colours present in some blackcurrant sweets can be separated by chromatography. The colours are water-soluble dyes. The diagrams show how the colours can be extracted from the sweets.



a Complete the empty boxes to name the pieces of apparatus. [3]

The apparatus below was used to carry out the chromatography.



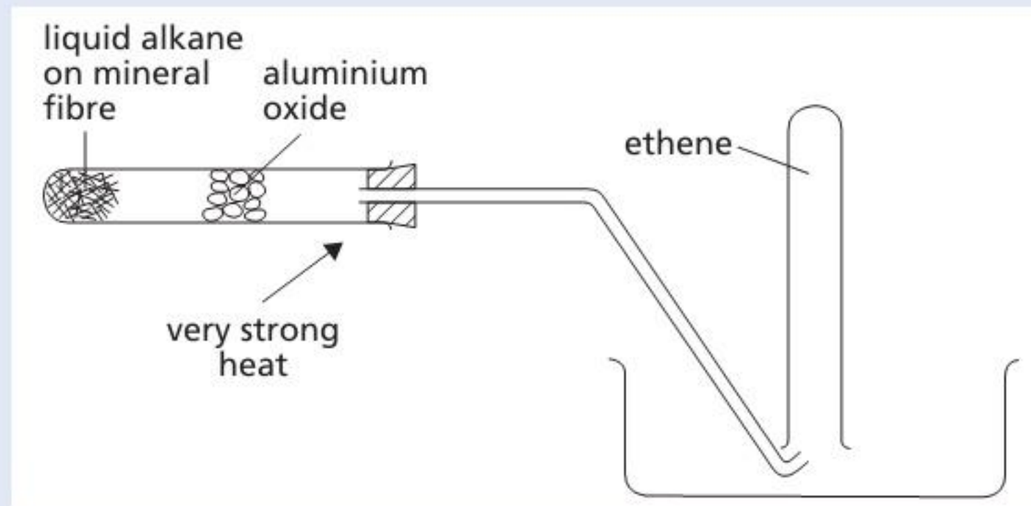
b (i) Name the solvent used. [1]
(ii) Label, with an arrow, the origin on the diagram. [1]

c Sketch the chromatogram you would expect if two different colours were present in the sweets. [1]

[Total: 6]

(Cambridge IGCSE Chemistry 0620 Paper 06 Q1 November 2008)

21 Ethene gas was formed by the cracking of a liquid alkane. The diagram shows the apparatus used.



a Identify two mistakes in the diagram. [2]

b Describe a test to show the presence of ethene.
test _____
result _____ [2]

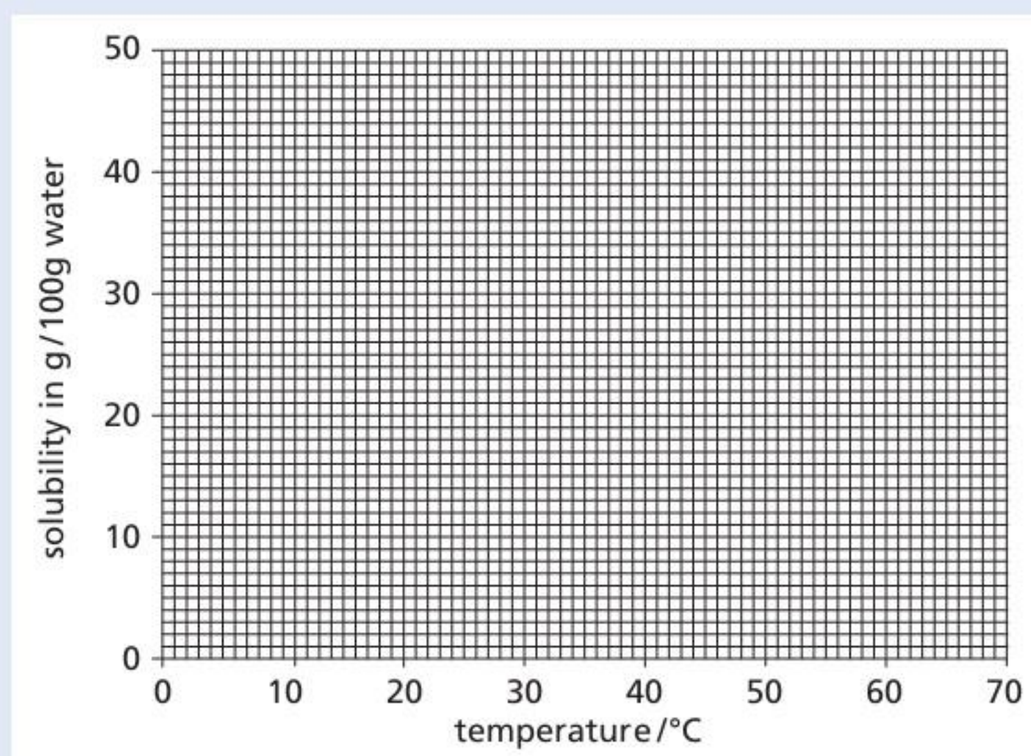
[Total: 4]

(Cambridge IGCSE Chemistry 0620 Paper 06 Q3 November 2008)

22 An experiment was carried out to determine the solubility of potassium chlorate at different temperatures. The solubility is the mass of potassium chlorate that dissolves in 100 g of water. The results obtained are shown in the table below.

Temperature/°C	0	10	20	30	40	50	60
Solubility in g/100 g water	12	17	20	24	29	34	40

a On the grid, draw a smooth line graph to show the solubility of potassium chlorate at different temperatures. [4]



b Use your graph to determine the solubility of potassium chlorate at 70 °C. Show clearly on the graph how you obtained your answer. [2]

c What would be the effect of cooling a saturated solution of potassium chlorate from 60 °C to 20 °C? [2]

[Total: 8]

(Cambridge IGCSE Chemistry 0620 Paper 06 Q6 November 2008)

23 A solution of magnesium sulfate can be made by reacting magnesium oxide with warm sulfuric acid.

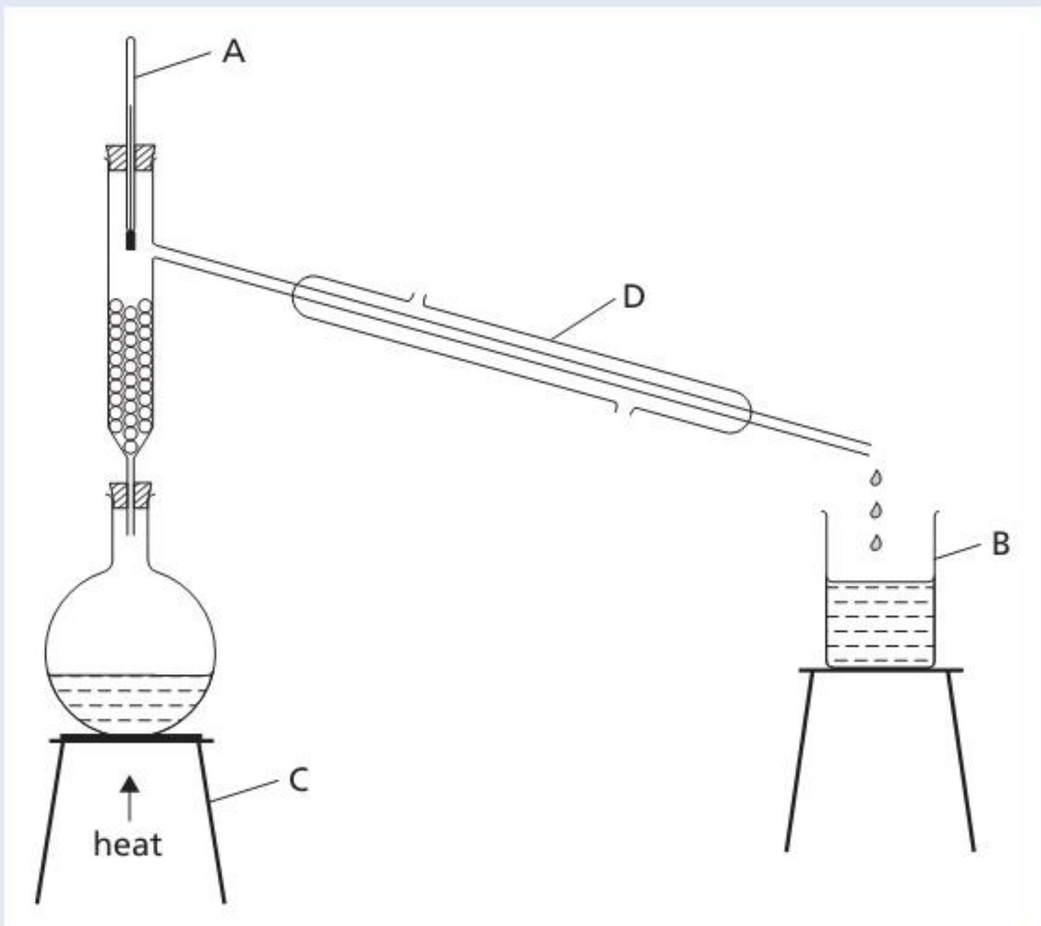
a Describe how you could make a solution of magnesium sulfate starting with magnesium oxide powder and dilute sulfuric acid. [3]

b Describe how you would obtain pure dry crystals of hydrated magnesium sulfate, $MgSO_4.7H_2O$, from the solution of magnesium sulfate in **a**. [3]

[Total: 6]

(Cambridge IGCSE Chemistry 0620 Paper 06 Q7 November 2008)

24 A mixture of ethanol and water can be separated by fractional distillation. The apparatus below can be used to carry out such a separation in the laboratory.



a Name each piece of apparatus. [3]

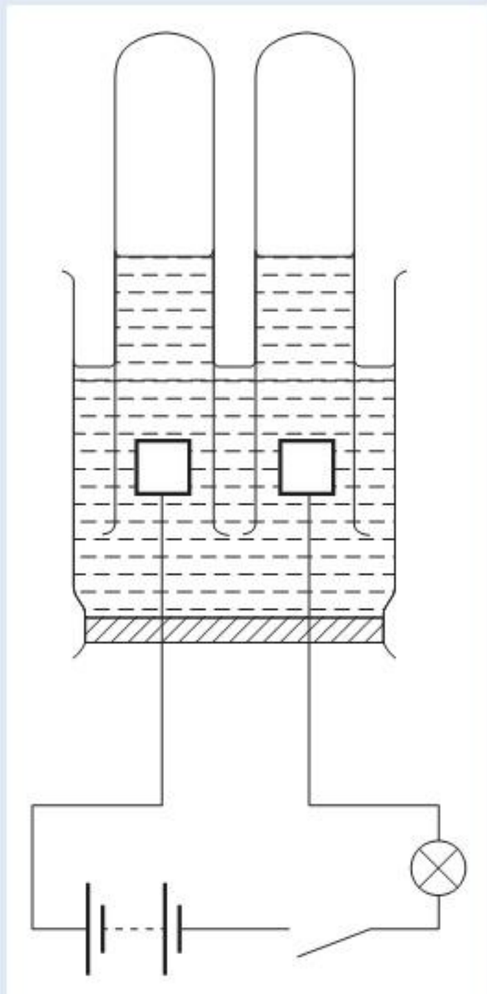
b What is the purpose of **D**? [1]

c How could the purity of the ethanol collected be checked? [1]

[Total: 5]

(Cambridge IGCSE Chemistry 0620 Paper 06 Q1 June 2007)

25 Concentrated hydrochloric acid can be electrolysed using the apparatus shown.



a Label the position of the electrodes on the diagram. [1]

b Give two observations when the circuit is switched on. [2]

c **(i)** Name the product at the positive electrode. [1]

(ii) State a test for this product and the result of the test.

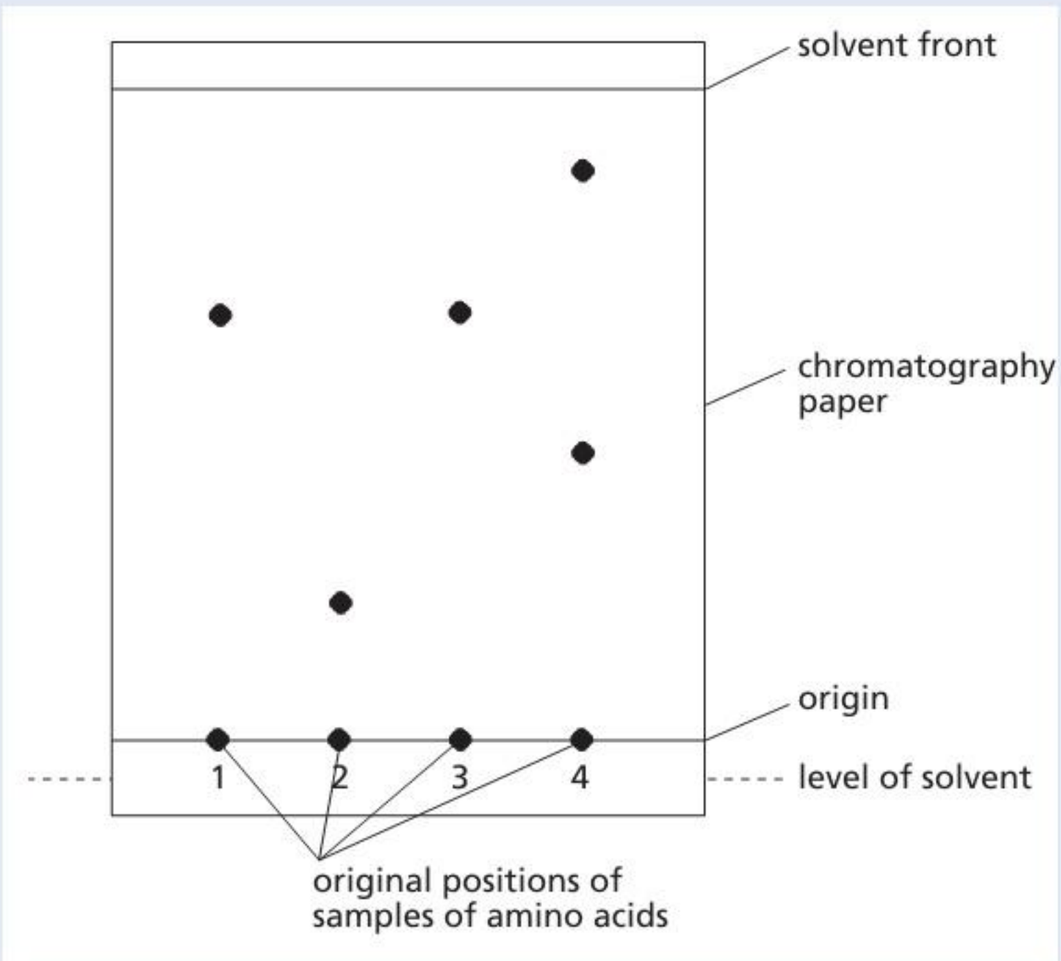
test _____

result _____ [2]

[Total: 6]

(Cambridge IGCSE Chemistry 0620 Paper 06 Q2 June 2007)

26 Chromatography can be used to identify amino acids from a sample of protein.
The diagram shows the chromatogram obtained when four samples of amino acids were analysed.
The paper was sprayed with ninhydrin.



a Why is the origin line drawn in pencil? [1]
b Which amino acids could possibly be the same? [1]
c Which amino acid sample contains more than one amino acid? Explain your answer.
sample _____
explanation _____ [2]
d Suggest why it is necessary to spray the chromatogram with ninhydrin. [1]

[Total: 5]

(Cambridge IGCSE Chemistry 0620 Paper 06 Q3 June 2007)

27 When cement powder is added to water a reaction takes place.
a Describe an experiment to show that this reaction is exothermic. [4]
b How could you show that the solution contains calcium ions? [2]

[Total: 6]

(Cambridge IGCSE Chemistry 0620 Paper 06 Q7 June 2007)

28 Seawater contains sodium chloride and other salts.
Plan an experiment to find the mass of salts in 1 dm^3 of seawater.
You will be provided with a small bottle of seawater.
You should include details of the method and any apparatus used.
(1 dm^3 = 1000 cm^3) [6]

[Total: 6]

(Cambridge IGCSE Chemistry 0620 Paper 61 Q6 October/November 2011)

29 Two different salts, **D** and **E**, were analysed.
D was an aqueous solution of iron(III) chloride and **E** was a solid.
The tests on the salts and some of the observations are in the following table.
Complete the observations in the table.

Tests	Observations
a (i) Appearance of solid **D**	..... [1]
(ii) Appearance of solid **E**	white crystals
Tests on solution D	
b The solution was divided into four equal portions in test-tubes, and the following tests carried out.	
(i) Dilute hydrochloric acid was added to the first portion of the solution and then aqueous barium chloride.	..... [1]
(ii) Dilute nitric acid was added to the second portion and then aqueous silver nitrate.	..... [2]
(iii) An excess of aqueous sodium hydroxide was added to the third portion of the solution.	..... [2]
(iv) An excess of aqueous ammonia was added to the fourth portion.	..... [1]
Test on solid E	
c (i) Solid **E** was heated in a test-tube. The gas given off was tested.	limewater turned milky
(ii) Dilute nitric acid was added to solid **E** in a test-tube.	rapid effervescence, limewater turned milky

d Identify the gas given off in tests **(c)(i)** and **(c)(ii)**. [1]

e What conclusions can you draw about solid **E**? [2]

[Total: 10]

(Cambridge IGCSE Chemistry 0620 Paper 63 Q5 October/November 2011)

30 Three different liquids **P**, **Q** and **R** were analysed. **P** was an aqueous solution of sulfuric acid. The tests on the liquids and some of the observations are in the following table. Complete the observations in the table.

Tests	Observations
a (i) Appearance of the liquids	**P** ..... [2] **Q** colourless, smell of vinegar **R** colourless, no smell
(ii) The pH of the liquids was tested using Universal indicator paper.	**P** ..... [2] **Q** pH 5 **R** pH 7
b A piece of magnesium ribbon was added to a little of each liquid. The gas given off by liquid **P** was tested.	**P** ..... [2] **Q** slow effervescence **R** no reaction
c To a little of liquid **P**, hydrochloric acid and aqueous barium chloride were added.	..... [2]
d Liquid **R** was heated to boiling in a test-tube. A thermometer was used to record the constant temperature of the vapour produced.	temperature = 100 °C

e What conclusions can you draw about liquid **Q**? [2]

f Identify liquid **R**. [1]

[Total: 9]

(Cambridge IGCSE Chemistry 0620 Paper 61 Q5 October/November 2011)

Theory

Papers 21 and 22 are Core theory papers; Papers 31, 32 and 33 are Extended theory papers.

Do not write in this book. Where necessary copy drawings, tables or sentences.

Physical

1 The diagram below shows the elements in a period of the Periodic Table.

Li	Be	B	C	N	O	F	Ne

a To which period of the Periodic Table do these elements belong? [1]

b Answer these questions using only the elements shown in the diagram. Each element can be used once, more than once or not at all.
Write down the symbol for the element which
(i) has six electrons in its outer shell.
(ii) is a halogen.
(iii) is a metal which reacts rapidly with cold water.
(iv) has two forms, graphite and diamond.
(v) is in Group II of the Periodic Table.
(vi) makes up about 80% of the air. [6]

c Complete the following sentence using words from the list below.

atoms **electrons** **molecules** **neutrons** **protons**

The _____ of the elements in the Periodic Table are arranged in order of increasing number of _____ . [2]

[Total: 9]

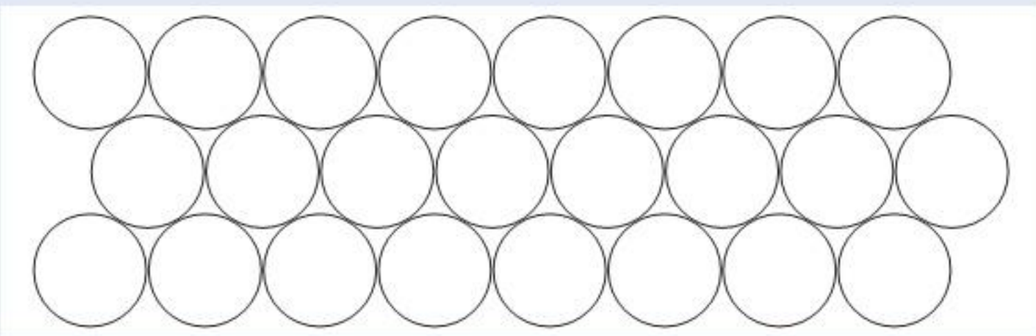
(Cambridge IGCSE Chemistry 0620 Paper 21 Q1 November 2010)

2 About 4000 years ago the Bronze Age started in Britain. Bronze is an alloy of copper and tin.

a (i) Suggest a reason why a bronze axe was better than a copper axe. [1]

(ii) Brass is another copper alloy. Name the other metal in brass. [1]

b The diagram shows the arrangement of particles in a pure metal.



(i) What is the name given to a regular arrangement of particles in a crystalline solid? [1]

(ii) Draw a diagram which shows the arrangement of particles in an alloy. [2]

(iii) Explain the term *malleable*. [1]

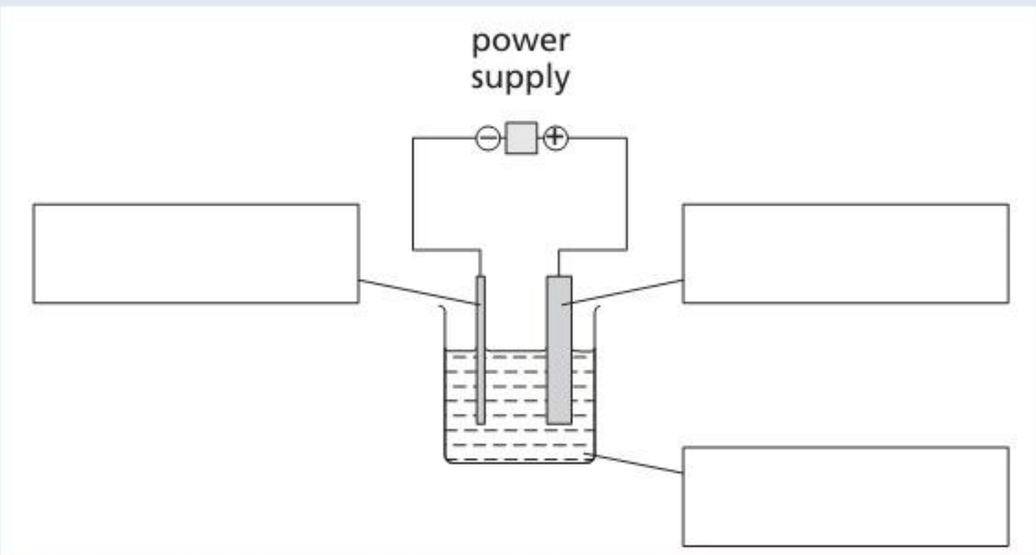
(iv) Why are metals malleable? [2]

c The common ore of tin is tin(IV) oxide and an ore of copper is malachite, $CuCO_3.Cu(OH)_2$.

(i) Write a word equation for the reduction of tin(IV) oxide by carbon. [1]

(ii) Malachite is heated to form copper oxide and two other chemicals. Name these chemicals. [2]

(iii) Copper oxide is reduced to copper which is then refined by electrolysis. Label the diagram of the apparatus which could be used to refine copper. [3]



(iv) Give **one** use of copper, other than making alloys. [1]

[Total: 15]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q2 November 2010)

3 Soluble salts can be made using a base and an acid.

a Complete this method of preparing dry crystals of the soluble salt cobalt(II) chloride-6-water from the insoluble base cobalt(II) carbonate.

Step 1 Add an excess of cobalt(II) carbonate to hot dilute hydrochloric acid.

Step 2

Step 3

Step 4 [4]

b 6.0 g of cobalt(II) carbonate was added to 40 cm^3 of hydrochloric acid, concentration 2.0 mol/dm^3. Calculate the maximum yield of cobalt(II) chloride-6-water and show that the cobalt(II) carbonate was in excess.

$$CoCO_3 + 2HCl \rightarrow CoCl_2 + CO_2 + H_2O$$

$$CoCl_2 + 6H_2O \rightarrow CoCl_2.6H_2O$$

Maximum yield

Number of moles of HCl used = _____

Number of moles of $CoCl_2$ formed = _____

Number of moles of $CoCl_2.6H_2O$ formed = _____

Mass of one mole of $CoCl_2.6H_2O$ = 238 g

Maximum yield of $CoCl_2.6H_2O$ = _____ g [4]

To show that cobalt(II) carbonate is in excess

Number of moles of HCl used = _____ (use value from above)

Mass of one mole of $CoCO_3$ = 119 g

Number of moles of $CoCO_3$ in 6.0 g of cobalt(II) carbonate = _____ [1]

Explain why cobalt(II) carbonate is in excess. [1]

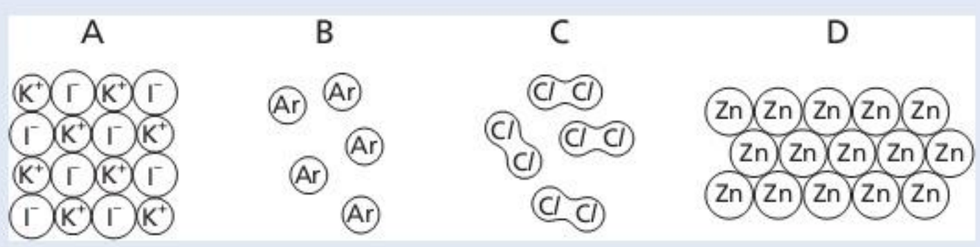
[Total: 10]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q8 November 2010)

4 The list describes five types of chemical structures.

giant covalent
giant ionic
metallic
simple atomic
simple molecular

a The diagrams below show four types of chemical structures.



(i) Use the list to match these structures with the diagrams.
Structure **A** is _____ [1]
Structure **B** is _____ [1]
Structure **C** is _____ [1]
Structure **D** is _____ [1]

(ii) Which **two** of the structures **A**, **B**, **C** and **D** have low melting points? [1]

b Sodium chloride is an ionic solid. Complete the following sentences using words from the list.

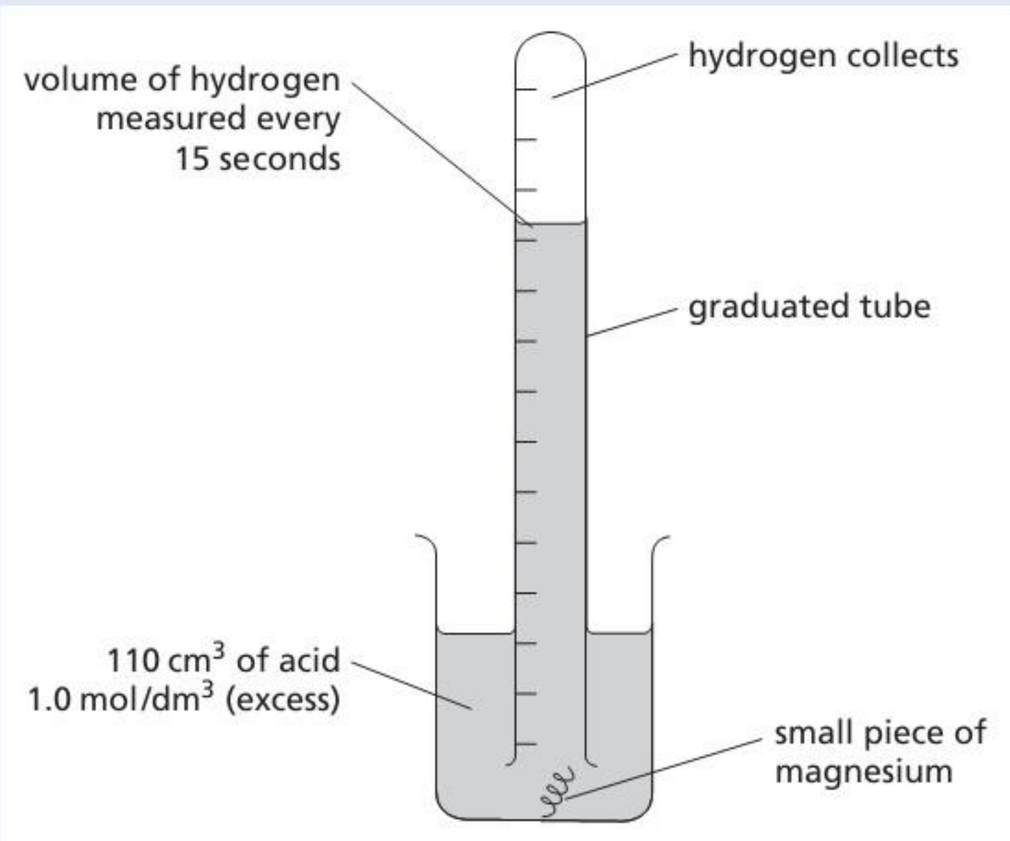
electrons ionic molecular molten solid

Sodium chloride does not conduct electricity when it is a _____ because the ions cannot move. When it is _____ sodium chloride does conduct electricity because the ions are free to move. [2]

[Total: 7]

(Cambridge IGCSE Chemistry 0620 Paper 22 Q2 June 2010)

5 A diagram of the apparatus which could be used to investigate the rate of reaction between magnesium and an excess of an acid is drawn below.

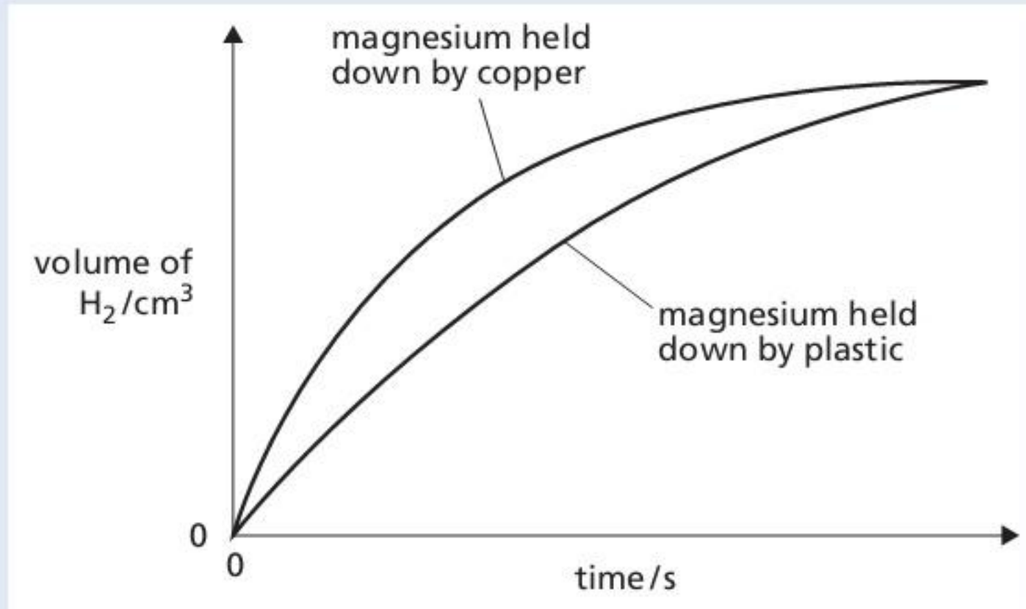


a The magnesium kept rising to the surface. In one experiment, this was prevented by twisting the magnesium around a piece of copper. In a second experiment, the magnesium was held down by a plastic net fastened to the beaker.

(i) Suggest a reason why magnesium, which is denser than water, floated to the surface. [1]

(ii) Iron, zinc and copper have similar densities. Why was copper a better choice than iron or zinc to weigh down the magnesium? [1]

b The only difference in the two experiments was the method used to hold down the magnesium. The results are shown.



(i) In which experiment did the magnesium react faster? [1]

(ii) Suggest a reason why the experiment chosen in **(i)** had the faster rate. [1]

c The experiment was repeated using 1.0 mol/dm³ propanoic acid instead of 1.0 mol/dm³ hydrochloric acid. Propanoic acid is a weak acid.

(i) How would the graph for propanoic acid **differ** from the graph for hydrochloric acid? [1]

(ii) How would the graph for propanoic acid be the **same** as the graph for hydrochloric acid? [1]

d Give **two** factors which would alter the rate of this reaction. For each factor explain why it alters the rate. [4]

[Total: 10]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q3 June 2010)

6 Iodine reacts with chlorine to form dark brown iodine monochloride.

$$I_2 + Cl_2 \rightarrow 2ICl$$

This reacts with more chlorine to give yellow iodine trichloride. There is an equilibrium between these iodine chlorides.

$$ICl(l) + Cl_2(g) \rightleftharpoons ICl_3(s)$$

a Explain what is meant by *equilibrium*. [2]

b When the equilibrium mixture is heated it becomes a darker brown colour. Is the reverse reaction endothermic or exothermic? Give a reason for your choice. [2]

c The pressure on the equilibrium mixture is decreased.

(i) How would this affect the position of equilibrium and why?
It would move to the _____ [1]
reason _____ [1]

(ii) Describe what you would observe. [1]

[Total: 7]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q6 June 2010)

7 Some grass is crushed and mixed with the solvent, propanone. The colour pigments are extracted to give a deep green solution.

a (i) Draw a labelled diagram to describe how you could show that there is more than one coloured pigment in the green solution. [3]

(ii) Given a pure sample of chlorophyll, how could you show that the green solution from the grass contained chlorophyll? [2]

b Explain the role of chlorophyll in green plants. [3]

[Total: 8]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q1 June 2009)

8 The following is a list of the electron distributions of atoms of unknown elements.

Element	Electron distribution
A	2,5
B	2,8,4
C	2,8,8,2
D	2,8,18,8
E	2,8,18,8,1
F	2,8,18,18,7

a Choose an element from the list for each of the following descriptions.

(i) It is a noble gas.

(ii) It is a soft metal with a low density.

(iii) It can form a covalent compound with element **A**.

(iv) It has a giant covalent structure similar to diamond.

(v) It can form a negative ion of the type X^{3-}. [5]

b Elements **C** and **F** can form an ionic compound.

(i) Draw a diagram that shows the formula of this compound, the charges on the ions and the arrangement of the valency electrons around the negative ion.
Use **o** to represent an electron from an atom of **C**.
Use **x** to represent an electron from an atom of **F**. [3]

(ii) Predict **two** properties of this compound. [2]

[Total: 10]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q3 June 2009)

9 Quantities of chemicals, expressed in moles, can be used to find the formula of a compound, to establish an equation and to determine reacting masses.

a A compound contains 72% magnesium and 28% nitrogen. What is its empirical formula? [2]

b A compound contains only aluminium and carbon. 0.03 moles of this compound reacted with excess water to form 0.12 moles of $Al(OH)_3$ and 0.09 moles of CH_4.
Write a balanced equation for this reaction. [2]

c 0.07 moles of silicon reacts with 25 g of bromine.

$$Si + 2Br_2 \rightarrow SiBr_4$$

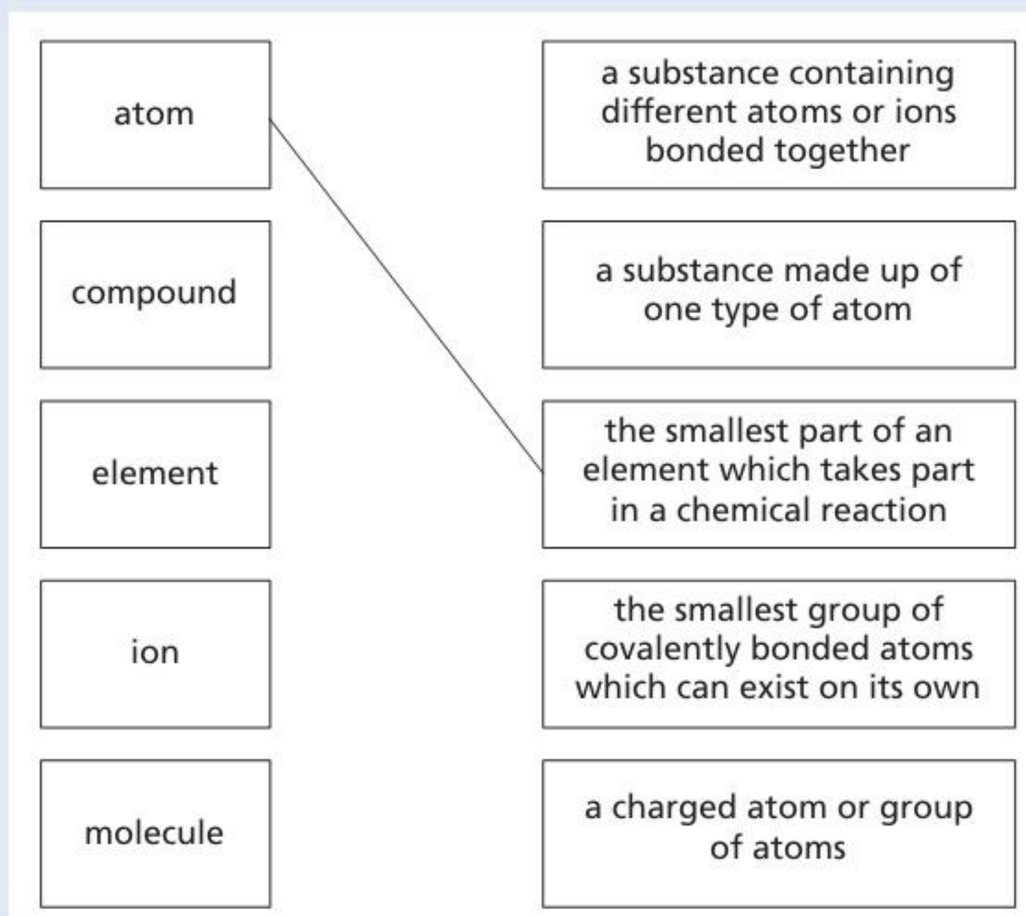
(i) Which one is the limiting reagent? Explain your choice. [3]

(ii) How many moles of $SiBr_4$ are formed? [1]

[Total: 8]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q9 June 2009)

10 a Link the terms in the boxes on the left with the definitions on the right. The first one has been done for you. [4]



b Which two of the following are mixtures? Tick two boxes.

air ☐

graphite ☐

sodium chloride ☐

steel ☐

[1]

c **(i)** Draw a labelled diagram to show the atomic structure of an atom of helium. In your diagram include the structure of the nucleus. [4]

(ii) State a use for helium. [1]

(iii) Which one of these statements about helium is correct? [1]

helium is in Period 2 of the Periodic Table ☐

helium is a liquid at room temperature ☐

helium is unreactive ☐

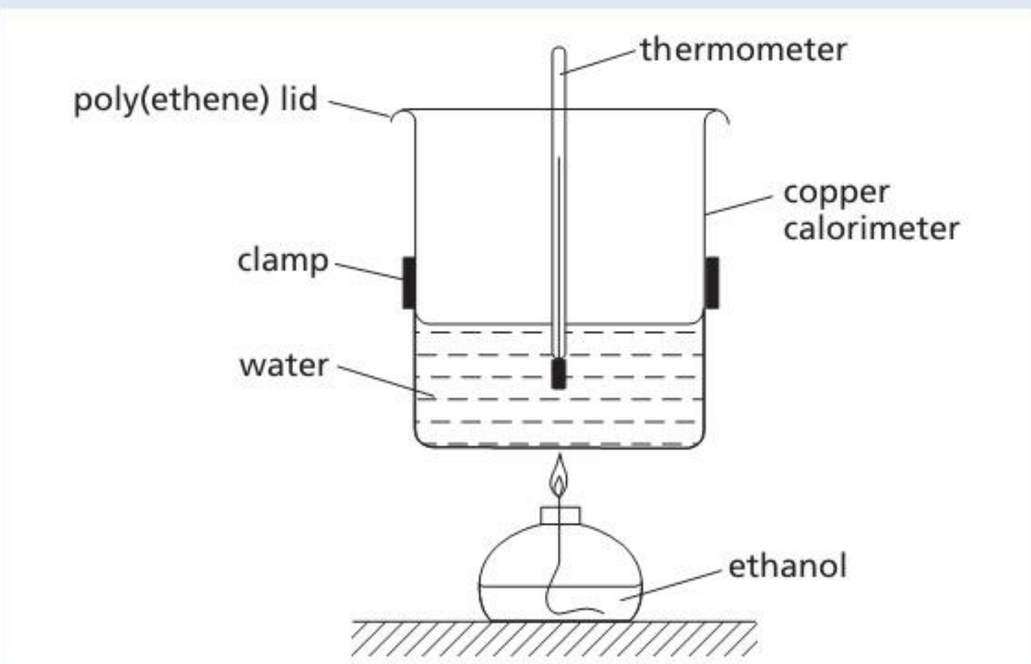
helium has an incomplete outer shell of electrons ☐

[1]

[Total: 11]

(Cambridge IGCSE Chemistry 0620 Paper 02 Q2 November 2009)

11 A student used the apparatus shown to calculate the energy released when ethanol burns.



a Draw the structure of ethanol showing all atoms and bonds. [1]

b The energy released by the burning ethanol raises the temperature of the water in the copper calorimeter.

(i) Which one of these words best describes the energy change when ethanol burns? Put a ring around the correct answer.

electrolytic electronic endothermic exothermic [1]

(ii) When 4.6 g of ethanol is burnt, 5.4 g of water is formed. Calculate the mass of water formed when 13.8 g of ethanol is burnt. [1]

(iii) Complete the equation for the combustion of ethanol.

$$C_2H_5OH + 3O_2 \rightarrow CO_2 + H_2O$$ [1]

c The calorimeter is made of copper. Copper is a transition metal. State two properties which distinguish transition metals from Group I metals. [2]

d When copper is left exposed to the air for some time, a coating of copper carbonate forms on its surface. The equation shows how copper carbonate reacts with hydrochloric acid.

$$CuCO_3(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + CO_2(g) + H_2O(l)$$

(i) Describe two observations that can be made as this reaction happens. [2]

(ii) State the meaning of the symbol (aq). [1]

e The calorimeter lid is made of poly(ethene). Complete these sentences about poly(ethene) using words from the list.

acids addition condensation ethane ethene monomers polymer

Poly(ethene) is a _____ formed by the _____ of ethene molecules.
In this reaction the ethene molecules can be described as _____ . [3]

[Total: 12]

(Cambridge IGCSE Chemistry 0620 Paper 02 Q3 November 2009)

12 Caesium is a metal in Group I of the Periodic Table.

a State two physical properties of caesium. [2]

b State the number of electrons in the outer shell of a caesium atom. [1]

c An isotope of caesium has a mass number of 133.

(i) What do you understand by the term *isotope*? [1]

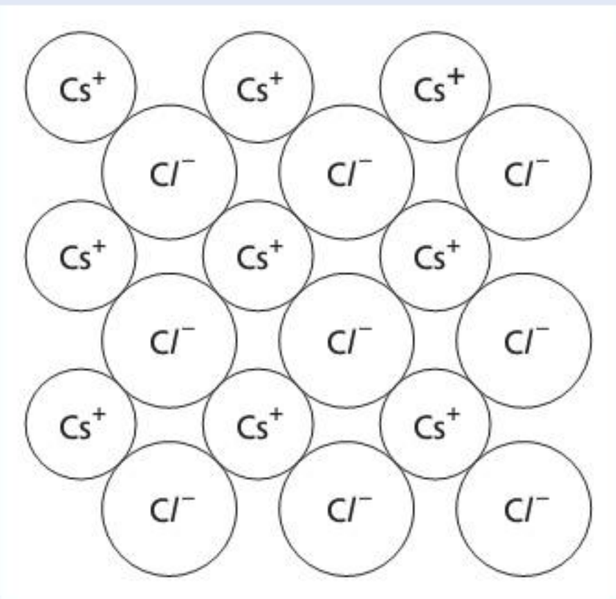
(ii) Calculate the number of neutrons in this isotope of caesium. [1]

d Complete the following table to estimate the boiling point of caesium and predict the reactivity of caesium with water.

Group I metal	Density/ g/cm³	Boiling point/°C	Reactivity with water
sodium	0.97	883	fizzes quickly, disappears gradually and does not burst into flame
potassium	0.86	760	fizzes very quickly, disappears quickly and bursts into flame with a little spitting
rubidium	1.53	686	fizzes extremely quickly, bursts into flame then spits violently and may explode
caesium	1.88		

[2]

e The diagram shows the structure of caesium chloride.



Use this diagram to work out the simplest formula for caesium chloride. [1]

f Caesium chloride dissolves in water to form a neutral solution. State the pH of a neutral solution. [1]

g Describe a test for chloride ions. [2]

[Total: 11]

(Cambridge IGCSE Chemistry 0620 Paper 02 Q4 November 2009)

13 The table shows the concentration of some ions present in seawater.

Name of ion	Formula of ion	Concentration of ion in g/dm³
bromide	Br^-	0.07
calcium	Ca^{2+}	0.4
chloride	Cl^-	19.1
magnesium	Mg^{2+}	1.2
potassium	K^+	0.3
sodium	Na^+	10.6
	SO_4^{2-}	0.8

a Which negative ion has the highest concentration in seawater? [1]

b State the name of the ion with the formula SO_4^{2-}. [1]

c Which two ions in the table are formed from Group I elements? [1]

d When seawater is evaporated a number of different compounds are formed. State the name of the compound which is present in the greatest quantity. [1]

e State the names of two ions in the table which move to the cathode when seawater is electrolysed. [2]

f When concentrated seawater is electrolysed, chlorine is formed at one of the electrodes.

(i) To which Period in the Periodic Table does chlorine belong? [1]

(ii) Draw the electronic structure of a chlorine molecule. Show only the outer electrons. [2]

g Drinking water can be obtained by purifying seawater.
Explain why distillation rather than filtration is used to purify seawater for drinking. [2]

[Total: 11]

(Cambridge IGCSE Chemistry 0620 Paper 02 Q3 June 2008)

14 For each of the following select an element from Period 4, potassium to krypton, that matches the description.

a It is a brown liquid at room temperature.

b It forms a compound with hydrogen having the formula XH_4.

c A metal that reacts violently with cold water.

d It has a complete outer energy level.

e It has oxidation states of 2 and 3 only.

f It can form an ion of the type X^-.
g One of its oxides is the catalyst in the Contact Process.

[Total: 7]
(Cambridge IGCSE Chemistry 0620 Paper 31 Q1 June 2008)

15 a Complete the table which gives the names, symbols, relative masses and relative charges of the three subatomic particles.

Name	Symbol	Relative mass	Relative charge
electron	e^-		
proton		1	
	n		0

[3]

b Use the information in the table to explain the following.
(i) Atoms contain charged particles but they are electrically neutral because they have no overall charge. [2]
(ii) Atoms can form positive ions. [2]
(iii) Atoms of the same element can have different masses. [2]
(iv) Scientists are certain that there are no undiscovered elements missing from the Periodic Table from hydrogen to lawrencium. [1]

[Total: 10]
(Cambridge IGCSE Chemistry 0620 Paper 31 Q2 June 2008)

16 Copper is purified by electrolysis.
a Complete the following.
The positive electrode (anode) is made from _____
The negative electrode (cathode) is made from _____
The electrolyte is aqueous _____ [3]
b Write an ionic equation for the reaction at the positive electrode (anode). [2]
c (i) Give **two** reasons why copper is used
in electric wiring, [2]
in cooking utensils. [2]
(ii) Give another use of copper. [1]

[Total: 10]
(Cambridge IGCSE Chemistry 0620 Paper 31 Q3 June 2008)

17 The following techniques are used to separate mixtures.

A simple distillation

B fractional distillation

C evaporation

D chromatography

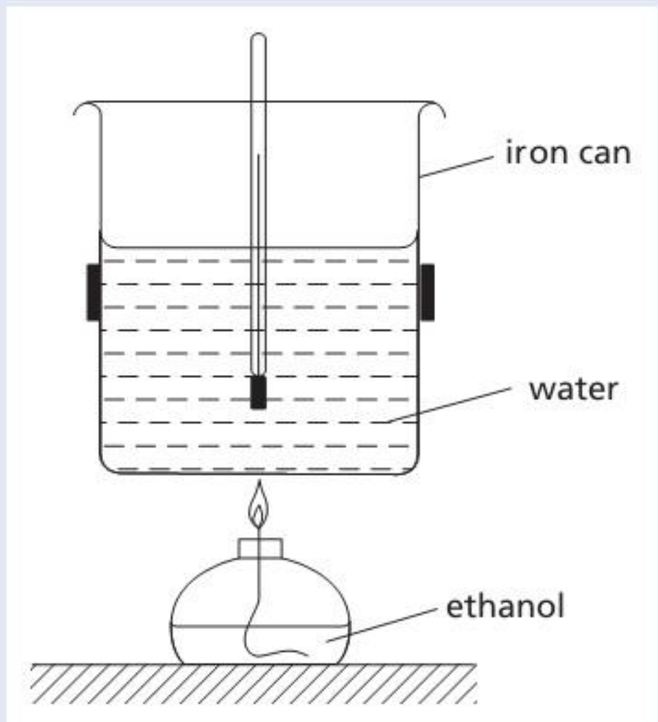
E filtration

F diffusion

From this list, choose the most suitable technique to separate the following.
a methane from a mixture of the gases, methane and ethane [1]
b water from aqueous magnesium sulfate [1]
c glycine from a mixture of the amino acids, glycine and lysine [1]
d iron filings from a mixture of iron filings and water [1]
e zinc sulfate crystals from aqueous zinc sulfate [1]
f hexane from a mixture of the liquids, hexane and octane [1]

[Total: 6]
(Cambridge IGCSE Chemistry 0620 Paper 31 Q1 June 2011)

18 The apparatus shown below can be used to measure the energy released when a liquid fuel is burnt. The amount of energy released is calculated from the increase in temperature of a known amount of water.



a **(i)** Explain how this experiment shows that the burning of ethanol is an exothermic reaction. [1]

(ii) Complete the word equation for the complete combustion of ethanol.

ethanol + oxygen → _____ + _____ [2]

b Ethanol is a fuel containing carbon. State the names of two other commonly used fuels containing carbon. [2]

c Give the formula of the functional group present in ethanol. [1]

d The can contains water. Describe a chemical test for water. [2]

e The iron can used in this experiment rusts easily.

(i) Describe a method which can be used to prevent iron from rusting. [1]

(ii) Rust contains hydrated iron(III) oxide. What do you understand by the term *hydrated*? [1]

(iii) Iron is a transition metal. State **two** properties which are typical of transition metals. [2]

[Total: 12]

(Cambridge IGCSE Chemistry 0620 Paper 02 Q5 November 2008)

19 There are three types of giant structure – ionic, metallic and macromolecular.

a Sodium nitride is an ionic compound. Draw a diagram that shows the formula of the compound, the charges on the ions and the arrangement of the valency electrons around the negative ion.
Use x to represent an electron from a sodium atom.
Use o to represent an electron from a nitrogen atom. [3]

b **(i)** Describe metallic bonding. [3]

(ii) Use the above ideas to explain why:
metals are good conductors of electricity, [1]
metals are malleable. [2]

c Silicon(IV) oxide has a macromolecular structure.

(i) **Describe** the structure of silicon(IV) oxide (a diagram is not acceptable). [3]

(ii) Diamond has a similar structure and consequently similar properties. Give **two** physical properties common to both diamond and silicon(IV) oxide. [2]

[Total: 14]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q2 November 2008)

20 The electrolysis of concentrated aqueous sodium chloride produces three commercially important chemicals: hydrogen, chlorine and sodium hydroxide.

a The ions present are $Na^+(aq)$, $H^+(aq)$,$Cl^-(aq)$ and $OH^-(aq)$.

(i) Complete the ionic equation for the reaction at the negative electrode (cathode).

_____ + _____ → H_2 [1]

(ii) Complete the ionic equation for the reaction at the positive electrode (anode).

_____ + _____ → Cl_2 [1]

(iii) Explain why the solution changes from sodium chloride to sodium hydroxide. [1]

b **(i)** Why does the water supply industry use chlorine? [1]

(ii) Name an important chemical that is made from hydrogen. [1]

(iii) How is sodium hydroxide used to make soap? [2]

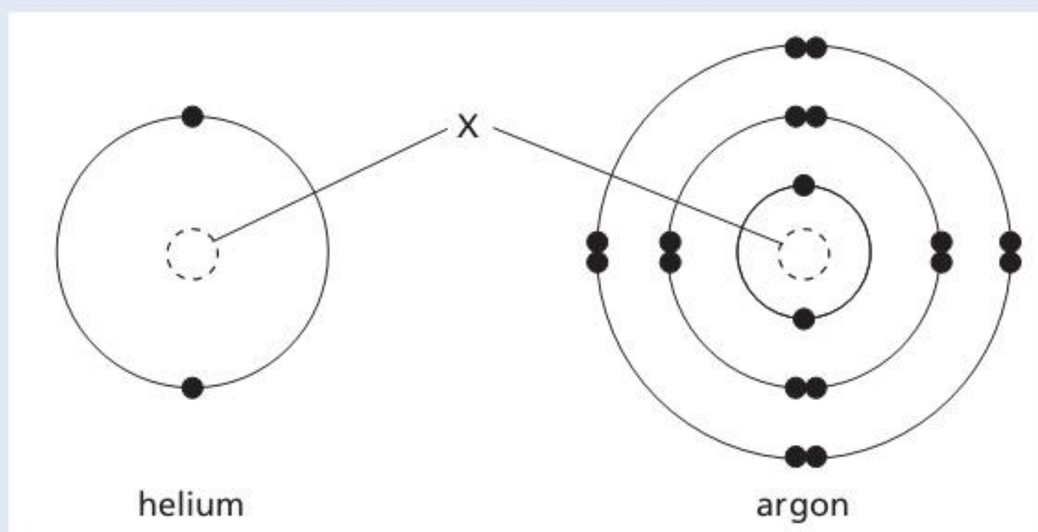
[Total: 7]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q5 November 2008)

21 Helium and argon are noble gases.

a State **one** use of helium. [1]

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



(i) State the name of the central part of the atom, labelled **X**. [1]

(ii) Which statement about helium and argon is correct? Tick **one** box.

Argon has an incomplete inner shell of electrons. ☐

An atom of argon has 16 electrons. ☐

Helium has a complete outer shell of electrons. ☐

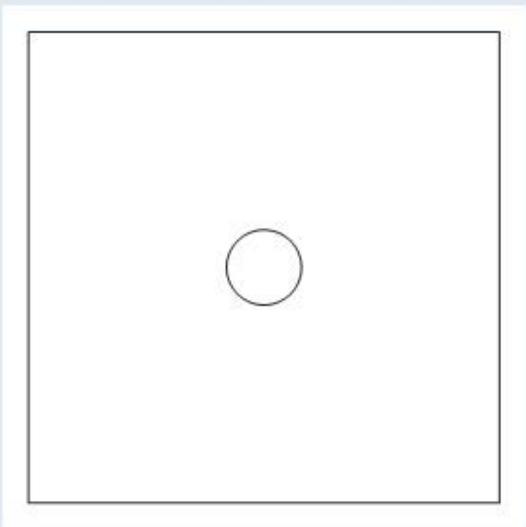
Helium has an incomplete outer shell of electrons. ☐

[1]

(iii) How many protons are there in an atom of argon? [1]

(iv) The symbol for a particular isotope of helium is written as $^{4}_{2}He$.
Write a similar symbol for the isotope of argon which has 16 neutrons. [1]

c Argon is a liquid at a temperature of –188 °C. Complete the diagram below to show how the atoms of argon are arranged at –188 °C.
◯ represents one atom of argon.



[2]

[Total: 7]
(Cambridge IGCSE Chemistry 0620 Paper 21 Q3 November 2010)

22 Hydrogen chloride can be made by burning hydrogen in chlorine.

a Complete the equation for this reaction.

H_2 + _____ → _____ HCl [2]

b Draw a dot and cross diagram for a molecule of hydrogen chloride. Show all the electrons.
Use **o** for an electron from a hydrogen atom.
Use **x** for an electron from a chlorine atom. [2]

c Hydrochloric acid is formed when hydrogen chloride gas dissolves in water. Suggest the pH of hydrochloric acid. Put a ring around the correct answer.

pH 1 **pH 7** **pH 9** **pH 13** [1]

d Complete the equation for the reaction of hydrochloric acid with zinc.

zinc + hydrochloric acid → zinc chloride + _____ [1]

e Describe how dry crystals of zinc chloride can be obtained from a solution of zinc chloride. [2]

f A student electrolysed molten zinc chloride. State the name of the product formed at:

(i) the anode, [1]

(ii) the cathode. [1]

[Total: 10]
(Cambridge IGCSE Chemistry 0620 Paper 02 Q7 June 2009)

23 Sulfuric acid is a typical strong acid.

a Change the equations given into a different format.

(i) $Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$
Change into a word equation. [1]

(ii) lithium oxide + sulfuric acid → lithium sulfate + water
Change into a symbol equation. [2]

(iii) $CuO + 2H^+ \rightarrow Cu^{2+} + H_2O$
Change the ionic equation into a symbol equation. [2]

(iv) $Na_2CO_3 + H_2SO_4 \rightarrow Na_2SO_4 + CO_2 + H_2O$
Change into a word equation. [1]

b When sulfuric acid dissolves in water, the following reaction occurs.

$$H_2SO_4 + H_2O \rightarrow HSO_4^- + H_3O^+$$

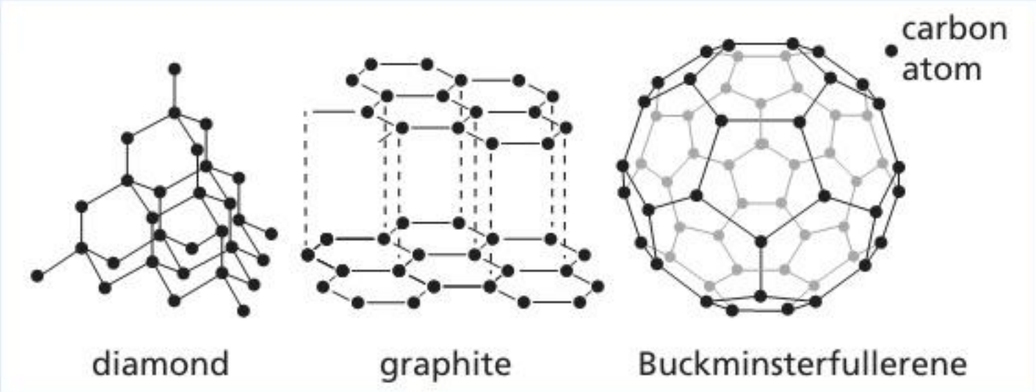
Explain why water is behaving as a base in this reaction. [2]

c Sulfuric acid is a strong acid, ethanoic acid is a weak acid.
Explain the difference between a strong acid and a weak acid. [2]

[Total: 10]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q4 June 2008)

24 Three forms of carbon are diamond, graphite and Buckminsterfullerene.



a **(i)** State **one** difference in structure between Buckminsterfullerene and diamond. [1]
(ii) State **two** differences in structure between graphite and diamond. [2]

b State the type of bonding between the carbon atoms in diamond. [1]

c Suggest why graphite is used as a lubricant. Refer to the layers in your answer. [1]

d State **one** use for diamond. [1]

e Coal is a fuel containing carbon. When coal is burnt, carbon dioxide is produced. Explain how the increase in carbon dioxide concentration in the atmosphere affects the world's climate. [2]

f Coal also contains small amounts of sulfur. Explain how burning coal leads to acid rain. [2]

g Methane is a fuel.
(i) Which one of the following is a natural source of methane?
Tick **one** box.

waste gases from respiration in plants ☐

waste gases from digestion in animals ☐

gases from photosynthesis in plants ☐

gases from forest fires ☐

[1]

(ii) Draw a diagram to show the arrangement of the electrons in a molecule of methane, CH_4.
Use
• for an electron from a carbon atom
x for an electron from a hydrogen atom. [1]

(iii) Methane belongs to the alkane homologous series.
Name **one** other alkane. [1]

[Total: 13]

(Cambridge IGCSE Chemistry 0620 Paper 22 Q7 June 2010)

Inorganic

25 Water is an important raw material in industry.

a State **one** use of water in industry. [1]

b Describe a chemical test for water.
test _____
result _____ [2]

c A small piece of potassium was placed in a beaker of water.
The equation for the reaction is

$$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$$

(i) Describe a test for the gas given off in this reaction.
test _____
result _____ [2]

(ii) What is the most likely pH of the solution in the beaker when the reaction is complete?
Put a ring around the correct answer.

pH2 **pH6** **pH7** **pH8** **pH12** [1]

d Water is formed when propane burns.

(i) Complete the equation for this reaction:

$C_3H_8 + 5O_2 \rightarrow$ _____ CO_2 + _____ H_2O [2]

(ii) Which of the following best describes this reaction?
Put a ring around the correct answer.

carbonisation **combustion**
dehydration **hydrogenation** [1]

(iii) When 4.4 g of propane are burnt in excess oxygen, 7.2 g of water are formed.
Calculate the mass of water formed when 22 g of propane are burnt. [1]

[Total: 10]

(Cambridge IGCSE Chemistry 0620 Paper 22 Q3 June 2010)

26 a Choose from the list of compounds to answer questions **(i)** to **(v)**.

calcium carbonate carbon dioxide
hydrogen chloride
iron(III) oxide lead(II) bromide
methane sodium hydroxide

Each compound can be used once, more than once or not at all.
Name the compound which
(i) is a transition metal compound, [1]
(ii) produces brown fumes at the anode when electrolysed, [1]
(iii) is used to manufacture lime, [1]
(iv) dissolves in water to form an alkaline solution, [1]
(v) is the main constituent of natural gas. [1]

b At a high temperature iron(III) oxide is reduced by carbon.

$$Fe_2O_3 + 3C \rightarrow 2Fe + 3CO$$

(i) Explain how the equation shows that iron(III) oxide is reduced by carbon. [1]
(ii) Complete these sentences about the extraction of iron using words from the list.

bauxite blast converter haematite
lime limestone sand slag

Iron is extracted from _____ by mixing the ore with coke and _____ in a _____ furnace.
The iron ore is reduced to iron and impurities in the ore react with calcium oxide to form _____. [4]

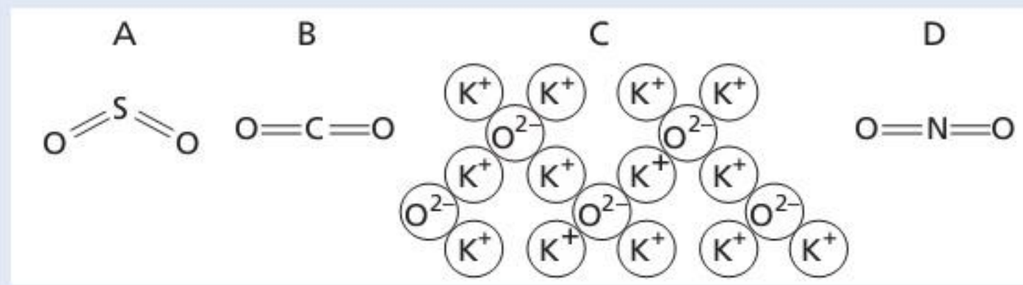
[Total: 10]

(Cambridge IGCSE Chemistry 0620 Paper 02 Q1 June 2009)

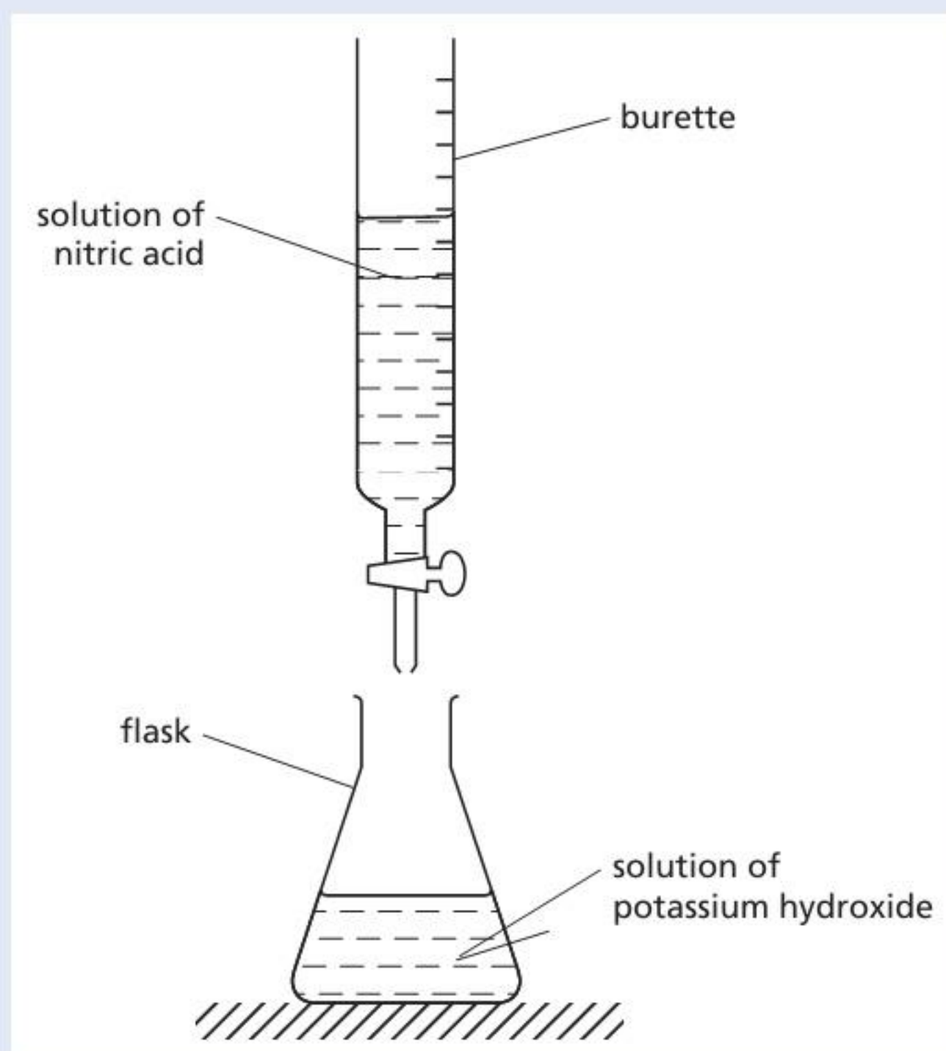
27 The list shows some non-metallic elements.

bromine
carbon
fluorine
krypton
nitrogen
oxygen

a Which **two** elements in the list are in the same Group of the Periodic Table? [1]
b Which element in the list has the highest proton number? [1]
c Which **two** of these elements make up most of the air? [1]
d Bromine and fluorine form a compound with the formula BrF_5. Calculate the relative molecular mass of BrF_5. [1]
e The diagram shows the structure of some compounds containing oxygen.



(i) What type of oxide is compound **C**? [1]
(ii) Compound **A** is an atmospheric pollutant. Describe the source of compound **A** and state its effect on the environment. [2]
(iii) In the presence of air, compound **D** reacts with water to form nitric acid.
A student used the apparatus shown to add an aqueous solution of nitric acid to an aqueous solution of potassium hydroxide. He added the acid until it was in excess.



Describe how the pH of the solution in the flask changes as the nitric acid is added until the acid is in excess. [3]

(iv) Describe how you can measure this pH change. [1]

(v) The equation for the reaction is

$$KOH + HNO_3 \rightarrow KNO_3 + H_2O$$

State the name of the salt formed in this reaction. [1]

[Total: 12]

(Cambridge IGCSE Chemistry 0620 Paper 02 Q1 November 2009)

28 A solution of calcium hydroxide in water is alkaline.

a Which **one** of the pH values below is alkaline? Put a ring around the correct answer.

pH 3 **pH 6** **pH 7** **pH 11** [1]

b Which of the following is the common name for calcium hydroxide?
Tick **one** box.

cement ☐

limestone ☐

quicklime ☐

slaked lime ☐

[1]

c Some farmers use calcium hydroxide to control soil acidity.

(i) Why is it important to control soil acidity? [1]

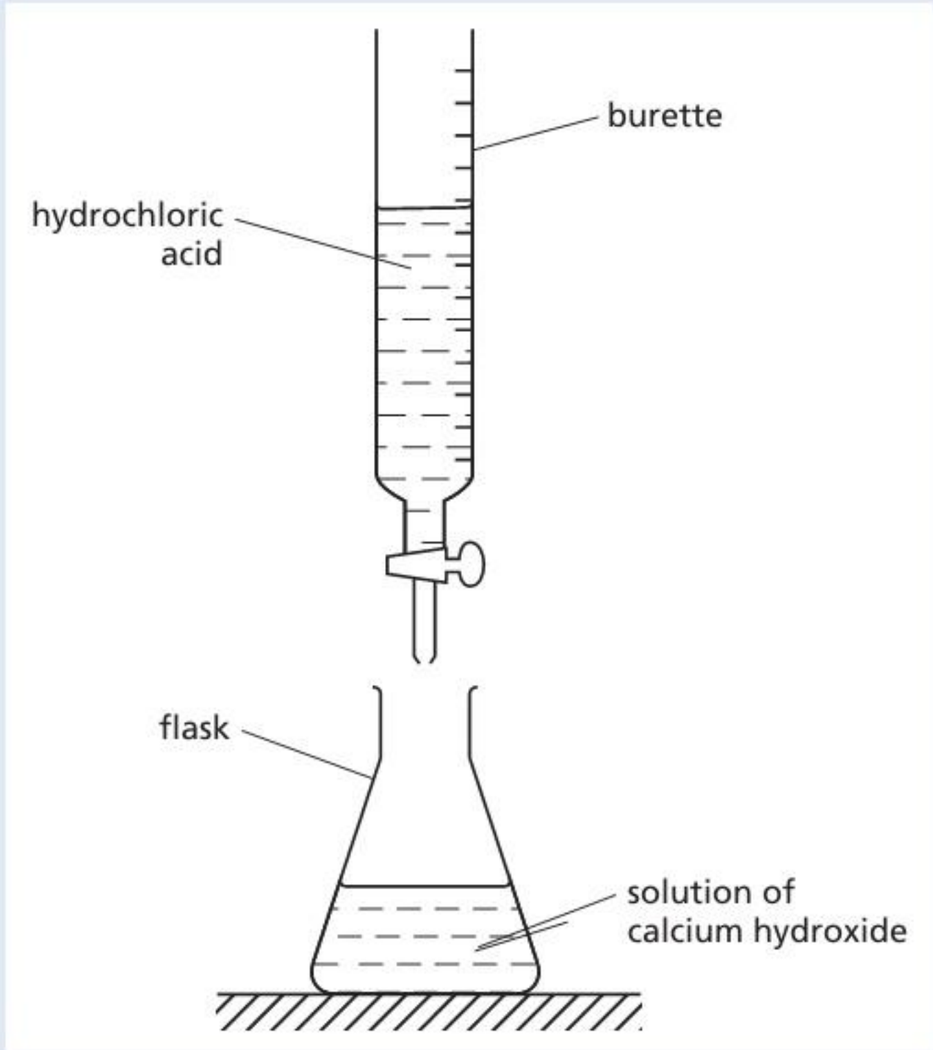
(ii) Acid rain can cause soil to become acidic. Describe how acid rain is formed. [3]

d Calcium hydroxide reacts with hydrochloric acid.

calcium hydroxide + hydrochloric acid → calcium chloride + water

(i) State the name of this type of chemical reaction. [1]

(ii) A dilute solution of calcium hydroxide can be titrated with hydrochloric acid using the apparatus shown.



Describe how you would carry out this titration. [3]

[Total: 10]

(Cambridge IGCSE Chemistry 0620 Paper 21 Q5 November 2010)

29 Ammonia is an important industrial chemical.

a (i) Give the electron structure of an atom of nitrogen. [1]

(ii) Use this electronic structure, rather than the valency of nitrogen, to explain why the formula of ammonia is NH_3 not NH_4. [2]

b Ammonia is made by the Haber Process.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

forward reaction is exothermic

The percentage of ammonia in the equilibrium mixture varies with conditions.

Pressure/atmospheres	100	200	300	400
% ammonia at 300 °C	45	65	72	78
% ammonia at 500 °C	9	18	25	31

The conditions actually used are 200 atmospheres, 450 °C and an iron catalyst.

(i) The original catalyst was platinum. Suggest a reason why it was changed to iron. [1]

(ii) Explain why the highest pressure gives the highest percentage of ammonia in the equilibrium mixture. [2]

(iii) What happens to the unreacted nitrogen and hydrogen? [1]

(iv) State **one** advantage and **one** disadvantage of using a lower temperature. [2]

[Total: 9]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q4 November 2010)

30 Lead is a grey metal.

a State **two** physical properties which are characteristic of metals. [2]

b To which Group in the Periodic Table does lead belong? [1]

c An isotope of lead has the mass number 208. Complete the table to show the number of subatomic particles in an atom of this isotope of lead. Use the Periodic Table to help you.

Type of particle	Number of particles
electrons	
protons	
neutrons	

[3]

d When lead is heated in oxygen, lead(II) oxide is formed. Write a word equation for this reaction. [1]

e When lead(II) oxide is heated with carbon, lead and carbon monoxide are formed.

$$PbO + C \rightarrow Pb + CO$$

(i) Which substance becomes oxidised during this reaction? [1]

(ii) Carbon monoxide is a covalent compound. Which one of these statements about carbon monoxide is correct? Tick **one** box.

It is a solid with a high melting point. ☐

It conducts electricity when it is a liquid. ☐

It is a gas at room temperature. ☐

It forms about 1% of the atmosphere. ☐

[1]

[Total: 9]

(Cambridge IGCSE Chemistry 0620 Paper 22 Q6 June 2010)

31 Choose an element which fits each of the following descriptions.

(i) It is a yellow solid which burns to form an acidic oxide. [1]

(ii) This element is a black solid which, when heated, forms a purple vapour. [1]

(iii) Most of its soluble salts are blue. [1]

(iv) It has a basic oxide of the type MO which is used to treat acidic soils. [1]

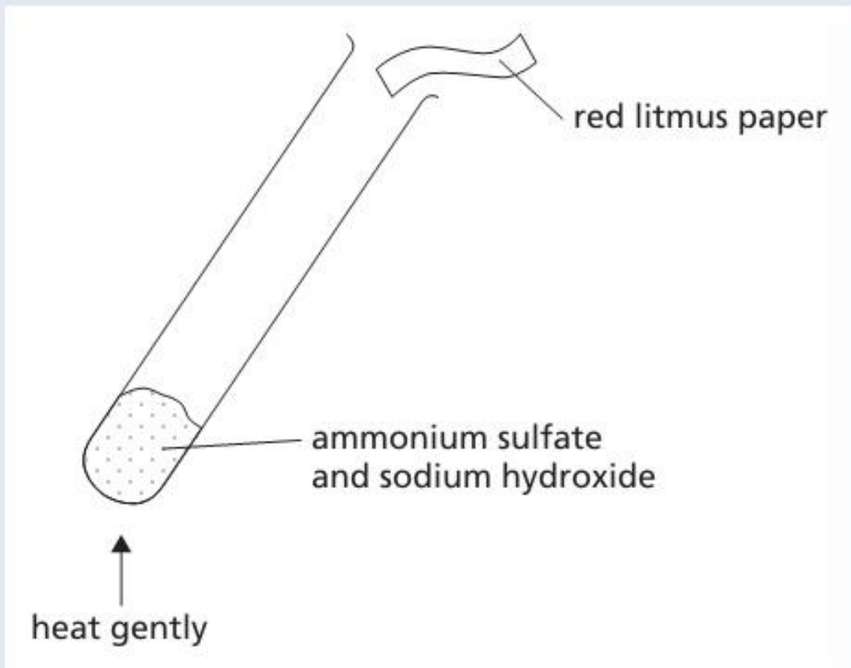
(v) It is an unreactive gas used to fill balloons. [1]

[Total: 5]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q1 June 2010)

32 This question is about some compounds of nitrogen.

A mixture of ammonium sulfate and sodium hydroxide was warmed in a test-tube. The gas was tested with moist red litmus paper.



a State the name of the gas released. [1]

b State the colour change of the litmus paper. [1]

c Complete the word equation for the reaction of ammonium carbonate with hydrochloric acid.

ammonium carbonate + hydrochloric acid → ___ + ___ + ___ [3]

d Ammonium salts such as ammonium nitrate, NH_4NO_3, and ammonium chloride, NH_4Cl, are used as fertilisers.

(i) Explain why farmers need to use fertilisers. [1]

(ii) Explain why ammonium nitrate is a better fertiliser than ammonium chloride. [1]

(iii) Calculate the relative formula mass of ammonium nitrate. [1]

e When ammonium nitrate is heated nitrogen(I) oxide is given off.
Nitrogen(I) oxide relights a glowing splint.
Name **one** other gas which relights a glowing splint. [1]

f State **one** harmful effect of nitrogen oxides on the environment. [1]

[Total: 10]

(Cambridge IGCSE Chemistry 0620 Paper 02 Q4 June 2009)

33 a The major gases in unpolluted air are 79% nitrogen and 20% oxygen.

(i) Name another gaseous element in unpolluted air. [1]

(ii) Name **two** compounds in unpolluted air. [2]

b Two common pollutants in air are carbon monoxide and the oxides of nitrogen.

(i) Name another pollutant in air. [1]

(ii) Describe how carbon monoxide is formed. [2]

(iii) How are the oxides of nitrogen formed? [2]

(iv) Explain how a catalytic converter reduces the emission of these two gases. [2]

[Total: 10]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q1 November 2009)

34 Oxides are classified as acidic, basic, neutral and amphoteric.

a Complete the table.

Type of oxide	pH of solution oxide	Example
acidic		
basic		
neutral		

[6]

b (i) Explain the term *amphoteric*. [1]

(ii) Name two reagents that are needed to show that an oxide is amphoteric. [2]

[Total: 9]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q2 November 2009)

35 a Sulfuric acid is made by the Contact process.

$$2SO_2 + O_2 \rightarrow 2SO_3$$

This is carried out in the presence of a catalyst at 450 °C and 2 atmospheres pressure.

(i) How is the sulfur dioxide made? [1]

(ii) Give another use of sulfur dioxide. [1]

(iii) Name the catalyst used. [1]

(iv) If the temperature is decreased to 300 °C, the yield of sulfur trioxide increases. Explain why this lower temperature is not used. [1]

(v) Sulfur trioxide is dissolved in concentrated sulfuric acid. This is added to water to make more sulfuric acid. Why is sulfur trioxide not added directly to water? [1]

b Sulfuric acid was first made in the Middle East by heating the mineral greenvitriol, $FeSO_4.7H_2O$. The gases formed were cooled.

$$\underset{\text{green crystals}}{FeSO_4.7H_2O(s)} \rightarrow \underset{\text{yellow powder}}{FeSO_4(s)} + 7H_2O(g)$$

$$2FeSO_4(s) \rightarrow Fe_2O_3(s) + SO_2(g) + SO_3(g)$$

On cooling

$$SO_3 + H_2O \rightarrow H_2SO_4 \text{ sulfuric acid}$$

$$SO_2 + H_2O \rightarrow H_2SO_3 \text{ sulfurous acid}$$

(i) How could you show that the first reaction is reversible? [2]

(ii) Sulfurous acid is a reductant. What would you see when acidified potassium manganate(VII) is added to a solution containing this acid? [2]

(iii) Suggest an explanation for why sulfurous acid in contact with air changes into sulfuric acid. [1]

c 9.12 g of anhydrous iron(II) sulfate was heated. Calculate the mass of iron(III) oxide formed and the volume of sulfur trioxide, at r.t.p., formed.

$$2FeSO_4(s) \rightarrow Fe_2O_3(s) + SO_2(g) + SO_3(g)$$

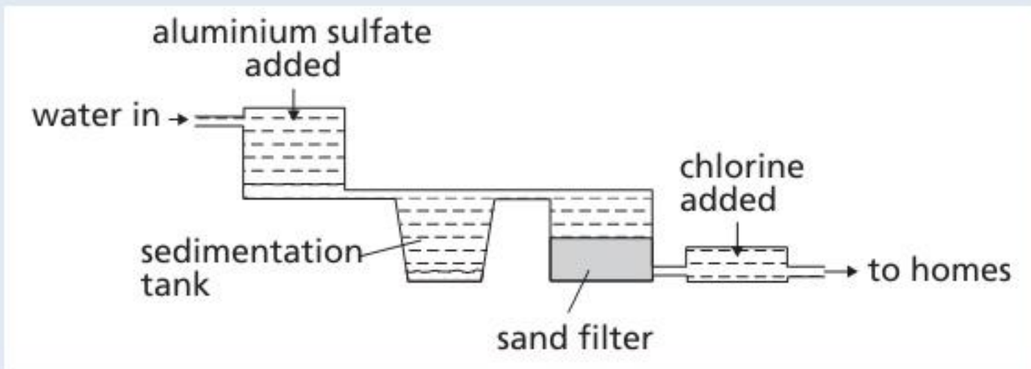
mass of one mole of $FeSO_4$ = 152 g
number of moles of $FeSO_4$ used = _____
number of moles of Fe_2O_3 formed = _____
mass of one mole of Fe_2O_3 = _____ g

mass of iron(III) oxide formed = _____ g
number of moles of SO_3 formed = _____
volume of sulfur trioxide formed = _____ dm^3
[6]

[Total: 16]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q6 November 2009)

36 The diagram shows a water treatment works.



a State one use of water in industry. [1]

b Explain how the sand filter helps purify the water. [2]

c The aluminium ions in aluminium sulfate cause clay particles to clump together.
Describe a test for aluminium ions.
test _____
result _____ [3]

d Why is chlorine added to the water? [1]

e Chlorine is in Group VII of the Periodic Table. When chlorine reacts with a solution of potassium bromide, the solution turns a reddish–brown colour.

(i) Write a word equation for this reaction. [2]

(ii) Explain why iodine does not react with a solution of potassium bromide. [1]

f When chlorine reacts with sodium to form sodium chloride, energy is released.

(i) State the name given to a reaction which releases energy. [1]

(ii) What type of bonding is present in sodium chloride? [1]

(iii) Explain what happens in terms of electron transfer when a sodium atom reacts with a chlorine atom. [2]

[Total: 14]

(Cambridge IGCSE Chemistry 0620 Paper 02 Q4 June 2008)

37 Selenium and sulfur are in Group VI. They have similar properties.

a One of the main uses of selenium is in photoelectric cells. These cells can change light into electrical energy.

(i) Name a process which can change light into chemical energy.

(ii) Name a device which can change chemical energy into electrical energy. [2]

b The electron distribution of a selenium atom is 2 + 8 + 18 + 6.

(i) Selenium forms an ionic compound with potassium. Draw a diagram which shows the formula of this ionic compound, the charges on the ions and the arrangement of the valency electrons around the negative ion.
Use o to represent an electron from an atom of potassium.
Use x to represent an electron from an atom of selenium. [3]

(ii) Draw a diagram showing the arrangement of the valency electrons in one molecule of the covalent compound selenium chloride.
Use x to represent an electron from an atom of selenium.
Use o to represent an electron from an atom of chlorine. [3]

(iii) Predict two differences in the physical properties of these two compounds. [2]

c The selenide ion reacts with water.

$$Se^{2-} + H_2O \rightarrow HSe^- + OH^-$$

What type of reagent is the selenide ion in this reaction? Give a reason for your choice. [3]

[Total: 13]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q2 June 2011)

38 Chlorine reacts with phosphorus to form phosphorus trichloride.

a Draw a diagram showing the arrangement of the valency electrons in one molecule of the covalent compound, phosphorus trichloride.
Use x to represent an electron from a phosphorus atom.

Use o to represent an electron from a chlorine atom. [2]

b Phosphorus trichloride reacts with water to form two acids.

(i) Balance the equation for this reaction.

PCl_3 + ___ H_2O → ___ HCl + H_3PO_3 [1]

(ii) Describe how you could show that phosphorus acid, H_3PO_3, is a weaker acid than hydrochloric acid. [3]

(iii) Two salts of phosphorus acid are its sodium salt, which is soluble in water, and its calcium salt which is insoluble in water. Suggest a method of preparation for each of these salts from aqueous phosphorus acid. Specify any other reagent needed and briefly outline the method.

sodium salt _____ [2]

calcium salt _____ [2]

[Total: 10]

(Cambridge IGCSE Chemistry 0620 Paper 33 Q7 June 2011)

39 Steel is an alloy made from impure iron.

a Both iron and steel rust. The formula for rust is $Fe_2O_3.2H_2O$. It is hydrated iron(III) oxide.

(i) Name the **two** substances that must be present for rusting to occur. [2]

(ii) Painting and coating with grease are two methods of preventing iron or steel from rusting. Give **two** other methods. [2]

b (i) Name a reagent that can reduce iron(III) oxide to iron. [1]

(ii) Write a symbol equation for the reduction of iron(III) oxide, Fe_2O_3, to iron. [2]

c (i) Calculate the mass of one mole of $Fe_2O_3.2H_2O$. [2]

(ii) Use your answer to **(i)** to calculate the percentage of iron in rust. [2]

d Iron from the blast furnace is impure. Two of the impurities are carbon and silicon. These are removed by blowing oxygen through the molten iron and adding calcium oxide.

(i) Explain how the addition of oxygen removes carbon. [1]

(ii) Explain how the addition of oxygen and calcium oxide removes silicon. [2]

[Total: 13]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q3 November 2008)

40 Ozone is a form of oxygen. Ozone is present in the upper atmosphere and it prevents dangerous solar radiation from reaching the Earth's surface. Some of the chemicals that diffuse into the upper atmosphere decompose ozone. Chemicals that have this effect are methane (CH_4), chloromethane (CH_3Cl) and an oxide of nitrogen (NO_2).

(i) Which of these three chemicals diffuses the most slowly? Give a reason for your choice. [2]

(ii) Chloromethane is formed when seaweed decomposes. Name the compounds in the environment from which seaweed might have obtained the following elements:

carbon; _____

hydrogen; _____

chlorine. _____ [3]

(iii) How can chloromethane be made from methane?

reagent _____

condition _____ [2]

(iv) The oxides of nitrogen are atmospheric pollutants. Describe how they are formed. [2]

(v) Complete the equation for the decomposition of ozone.

_____ O_3 → _____ [2]

[Total: 11]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q2 June 2010)

Organic

41 Ethene, C_2H_4, is manufactured by cracking petroleum fractions.

a (i) What do you understand by the term *petroleum fraction*? [1]

(ii) Complete the equation for the manufacture of ethene from dodecane, $C_{12}H_{26}$.

$C_{12}H_{26}$ → C_2H_4 + _____ [1]

b Two fractions obtained from the distillation of petroleum are refinery gas and gasoline. State **one** use of each of these fractions. [2]

c Ethene is an unsaturated hydrocarbon. What do you understand by the following terms?

unsaturated

hydrocarbon [2]

d Ethene is used to make ethanol.

(i) Which of these reactions is used to make ethanol from ethene? Tick **one** box.

- catalytic addition of steam ☐
- fermentation ☐
- oxidation using oxygen ☐
- reduction using hydrogen ☐

[1]

(ii) Draw the structure of ethanol showing all atoms and bonds. [2]

e Ethene is used to make poly(ethene). Complete the following sentences about this reaction. Use words from the list below.

additions carbohydrates catalysts monomers polymers

The ethene molecules which join to form poly(ethene) are the _____ .
The poly(ethene) molecules formed are _____ . [2]

[Total: 11]

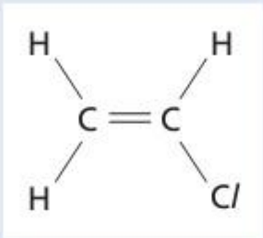
(Cambridge IGCSE Chemistry 0620 Paper 21 Q7 November 2010)

42 Monomers polymerise to form polymers or macromolecules.

a (i) Explain the term *polymerise*. [1]

(ii) There are two types of polymerisation – addition and condensation.
What is the difference between them? [2]

b An important monomer is chloroethene which has the structural formula shown below.



It is made by the following method.

$C_2H_4 + Cl_2 \rightarrow C_2H_4Cl_2$ dichloroethane

This is heated to make chloroethene.

$C_2H_4Cl_2 \rightarrow C_2H_3Cl + HCl$

(i) Ethene is made by cracking alkanes. Complete the equation for cracking dodecane.

$C_{12}H_{26} \rightarrow$ _____ $+ 2C_2H_4$ [1]

Another method of making dichloroethane is from ethane.

$C_2H_6 + 2Cl_2 \rightarrow C_2H_4Cl_2 + 2HCl$

(ii) Suggest a reason why the method using ethene is preferred. [1]

(iii) Describe an industrial method of making chlorine. [2]

(iv) Draw the structural formula of poly(chloroethene). Include three monomer units. [2]

[Total: 9]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q5 November 2010)

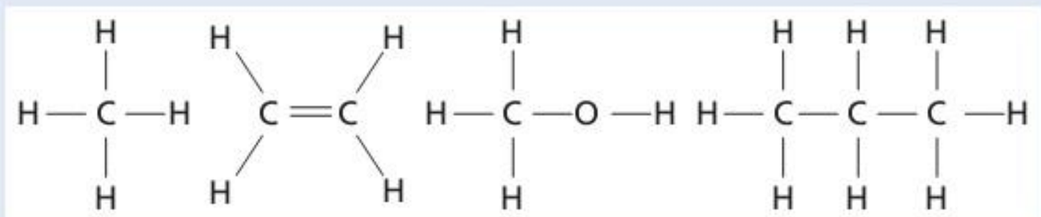
43 Petroleum is a mixture of hydrocarbons which can be separated into fractions such as petrol, paraffin and diesel.

a State the name of the process used to separate these fractions. [1]

b Name **two** other fractions which are obtained from petroleum. [2]

c Give **one** use for the paraffin fraction. [1]

d Many of the compounds obtained from petroleum are alkanes. Which **two** of the following structures are alkanes? [1]



e Use words from the list below to complete the following sentence.

ethane ethene hydrogen nitrogen oxygen reactive unreactive water

Alkanes such as _____ are generally _____ but they can be burnt in _____ to form carbon dioxide and _____ . [4]

f Alkanes are saturated hydrocarbons.
What do you understand by the terms

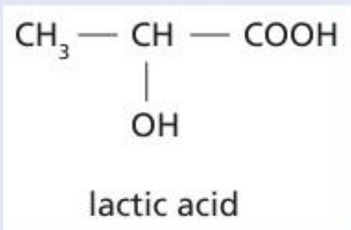
(i) saturated,

(ii) hydrocarbon? [2]

[Total: 11]

(Cambridge IGCSE Chemistry 0620 Paper 02 Q3 June 2009)

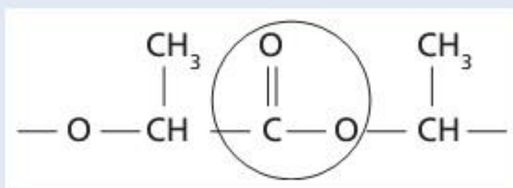
44 Lactic acid can be made from corn starch.



It polymerises to form the polymer, polylactic acid (PLA) which is biodegradable.

a Suggest **two** advantages that PLA has compared with a polymer made from petroleum. [2]

b The structure of PLA is given below.

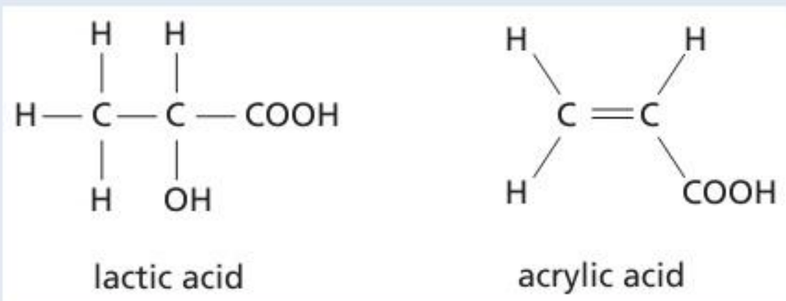


(i) What type of compound contains the group that is circled? [1]

(ii) Complete the following sentence.
Lactic acid molecules can form this group because they contain both an _____ group and an _____ group. [2]

(iii) Is the formation of PLA an addition or condensation polymerisation? Give a reason for your choice. [2]

c When lactic acid is heated, acrylic acid is formed.



(i) Complete the word equation for the action of heat on lactic acid.

lactic acid → _____ + _____ [1]

(ii) Describe a test that would distinguish between lactic acid and acrylic acid.

test result for lactic acid _____
result for acrylic acid _____ [3]

(iii) Describe a test, other than using an indicator, which would show that both chemicals contain an acid group.
test _____
result _____ [2]

[Total: 13]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q8 June 2009)

45 Butan-1-ol is used as a solvent for paints and varnishes, to make esters and as a fuel. Butan-1-ol can be manufactured from but-1-ene, which is made from petroleum.
Biobutanol is a fuel of the future. It can be made by the fermentation of almost any form of biomass – grain, straw, leaves etc.

a But-1-ene can be obtained from alkanes such as decane, $C_{10}H_{22}$, by cracking.

(i) Give the reaction conditions. [2]

(ii) Complete an equation for the cracking of decane, $C_{10}H_{22}$, to give but-1-ene.

$$C_{10}H_{22} \rightarrow _____$$ [2]

(iii) Name the reagent that reacts with but-1-ene to form butan-1-ol. [1]

b (i) Balance the equation for the complete combustion of butan-1-ol.

$$___\ C_4H_9OH + ___\ O_2 \rightarrow ___\ CO_2 + ___\ H_2O$$ [2]

(ii) Write a word equation for the preparation of the ester butyl methanoate. [2]

c The fermentation of biomass by bacteria produces a mixture of products which include biobutanol, propanol, hydrogen and propanoic acid.

(i) Draw the structural formula of propanol and of propanoic acid. Show all the bonds. [2]

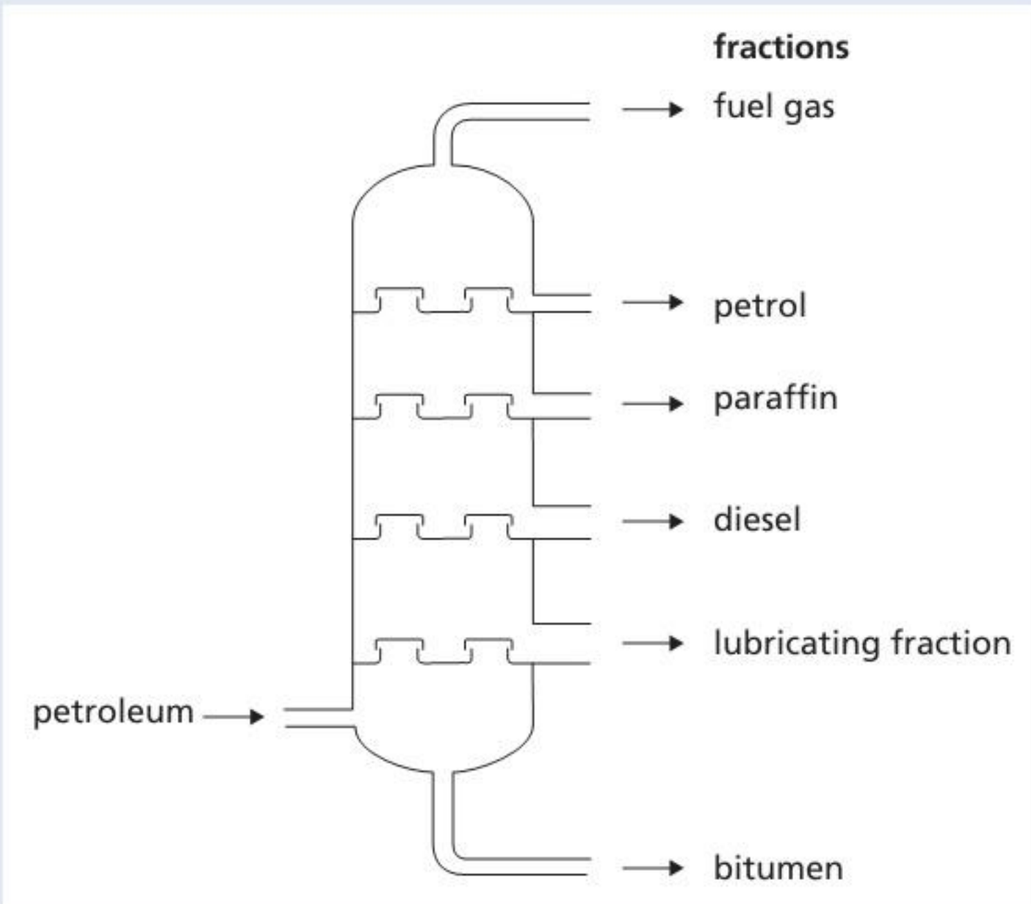
(ii) Why is it important to develop these fuels, such as biobutanol, as alternatives to petroleum? [1]

d How could you show that butanol made from petroleum and biobutanol are the same chemical? [1]

[Total: 13]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q7 November 2009)

46 Petroleum is separated into useful fractions by distillation.



a **(i)** What do you understand by the term *fraction*? [1]

(ii) Which fraction has the lowest boiling point? [1]

(iii) Describe how distillation is used to separate these fractions. [2]

(iv) State a use for
the paraffin fraction,
the bitumen fraction. [2]

b Ethene can be made by cracking certain hydrocarbon fractions.

(i) Explain what is meant by the term *cracking*. [1]

(ii) Complete the equation for the cracking of tetradecane, $C_{14}H_{30}$.

$$C_{14}H_{30} \rightarrow _____ + C_2H_4$$ [1]

c Ethanol is formed when steam reacts with ethene at high pressure and temperature. A catalyst of phosphoric acid is used.

$$\text{ethene} + \text{steam} \rightleftharpoons \text{ethanol}$$

(i) What is the function of the catalyst? [1]

(ii) What is the meaning of the symbol $\rightleftharpoons$? [1]

(iii) Ethanol is also formed when yeast grows in sugar solution. What is this process called?
Put a ring around the correct answer.

addition **combustion** **fermentation** **neutralisation** [1]

(iv) Phosphoric acid is a typical acid. State what you would observe when a solution of phosphoric acid is added to
blue litmus,
a solution of sodium carbonate. [2]

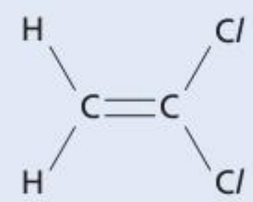
[Total: 13]

(Cambridge IGCSE Chemistry 0620 Paper 02 Q6 June 2008)

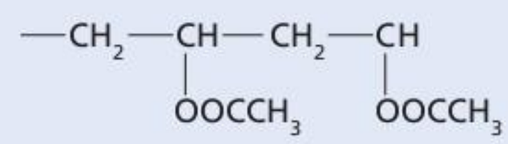
47 There are two types of polymerisation – addition and condensation.

a Explain the difference between them. [2]

b Poly(dichloroethene) is used to package food. Draw its structure. The structural formula of dichloroethene is shown below. [2]



c The polymer known as PVA is used in paints and adhesives. Its structural formula is shown below.



Deduce the structural formula of its monomer. [1]

d A condensation polymer can be made from the following monomers.

$HOOC(CH_2)_4COOH$ and $H_2N(CH_2)_6NH_2$

Draw the structural formula of this polymer. [3]

[Total: 8]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q8 June 2011)

48 The alkanes are generally unreactive. Their reactions include combustion, substitution and cracking.

a The complete combustion of an alkane gives carbon dioxide and water.

(i) $10\,cm^3$ of butane is mixed with $100\,cm^3$ of oxygen, which is an excess. The mixture is ignited. What is the volume of unreacted oxygen left and what is the volume of carbon dioxide formed?

$$C_4H_{10}(g) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$$

Volume of oxygen left = _____ cm^3
Volume of carbon dioxide formed = _____ cm^3 [2]

(ii) Why is the incomplete combustion of any alkane dangerous, particularly in an enclosed space? [2]

b The equation for a substitution reaction of butane is given below.

$$CH_3{-}CH_2{-}CH_2{-}CH_3 + Cl_2 \rightarrow CH_3{-}CH_2{-}CH_2{-}CH_2{-}Cl + HCl$$ [1]

(i) Name the organic product. [1]

(ii) This reaction does not need increased temperature or pressure. What is the essential reaction condition? [1]

(iii) Write a different equation for a substitution reaction between butane and chlorine. [1]

c Alkenes are more reactive and industrially more useful than alkanes.
They are made by cracking alkanes.

$$\underset{\text{heptane}}{C_7H_{16}} \rightarrow \underset{\text{propene}}{CH_3{-}CH{=}CH_2} + \underset{\text{but-1-ene}}{CH_3{-}CH_2{-}CH{=}CH_2} + H_2$$

(i) Draw the structural formula of the polymer poly(propene). [2]

(ii) Give the structural formula and name of the alcohol formed when but-1-ene reacts with steam. [2]

(iii) Deduce the structural formula of the product formed when propene reacts with hydrogen chloride. [1]

[Total: 12]

(Cambridge IGCSE Chemistry 0620 Paper 31 Q7 November 2008)

294 The Periodic Table of the elem

I	II							
						1 H Hydrogen 1		
7 Li Lithium 3	9 Be Beryllium 4							
23 Na Sodium 11	24 Mg Magnesium 12							
39 K Potassium 19	40 Ca Calcium 20	45 Sc Scandium 21	48 Ti Titanium 22	51 V Vanadium 23	52 Cr Chromium 24	55 Mn Manganese 25	56 Fe Iron 26	59 Co Cobalt 27
85 Rb Rubidium 37	88 Sr Strontium 38	89 Y Yttrium 39	91 Zr Zirconium 40	93 Nb Niobium 41	96 Mo Molybdenum 42	Tc Technetium 43	101 Ru Ruthenium 44	103 Rh Rhodium 45
133 Cs Caesium 55	137 Ba Barium 56	139 La Lanthanum 57 *	178 Hf Hafnium 72	181 Ta Tantalum 73	184 W Tungsten 74	186 Re Rhenium 75	190 Os Osmium 76	192 Ir Irodium 77
Fr Francium 87	226 Ra Radium 88	227 Ac Actinium 89 †						

*58-71 Lanthanoid series
†90-103 Actinoid series

140 Ce Cerium 58	141 Pr Praseodymium 59	144 Nd Neodymium 60	Pm Promethium 61	150 Sm Samarium 62
232 Th Thorium 90	Pa Protactinium 91	238 U Uranium 92	Np Neptunium 93	Pu Plutonium 94

Key

a X b	a = relative atomic mass X = atomic symbol b = proton (atomic) number

The volume of one mole of any gas is 2

nents

Group																		
							III	IV	V	VI	VII	0
												4 He Helium 2
							11 B Boron 5	12 C Carbon 6	14 N Nitrogen 7	16 O Oxygen 8	19 F Fluorine 9	20 Ne Neon 10
							27 Al Aluminium 13	28 Si Silicon 14	31 P Phosphorus 15	32 S Sulfur 16	35.5 Cl Chlorine 17	40 Ar Argon 18
	59 Ni Nickel 28	64 Cu Copper 29	65 Zn Zinc 30	70 Ga Gallium 31	73 Ge Germanium 32	75 As Arsenic 33	79 Se Selenium 34	80 Br Bromine 35	84 Kr Krypton 36			
	106 Pd Palladium 46	108 Ag Silver 47	112 Cd Cadmium 48	115 In Indium 49	119 Sn Tin 50	122 Sb Antimony 51	128 Te Tellurium 52	127 I Iodine 53	131 Xe Xenon 54			
	195 Pt Platinum 78	197 Au Gold 79	201 Hg Mercury 80	204 Tl Thallium 81	207 Pb Lead 82	209 Bi Bismuth 83	Po Polonium 84	At Astatine 85	Rn Radon 86			

152 Eu Europium 63	157 Gd Gadolinium 64	159 Tb Terbium 65	163 Dy Dysprosium 66	165 Ho Holmium 67	167 Er Erbium 68	169 Tm Thulium 69	173 Yb Ytterbium 70	175 Lu Lutetium 71
Am Americium 95	Cm Curium 96	Bk Berkelium 97	Cf Californium 98	Es Einsteinium 99	Fm Fermium 100	Md Mendelevium 101	No Nobelium 102	Lr Lawrecium 103

24 dm^3 at room temperature and pressure (r.t.p.).

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