

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:

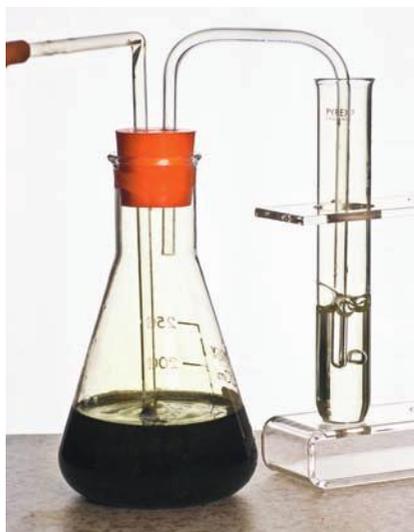
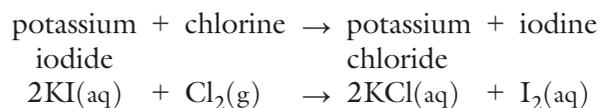
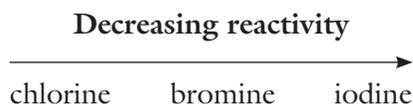


Figure 9.13 Iodine being displaced from potassium iodide solution as chlorine is bubbled through.

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



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

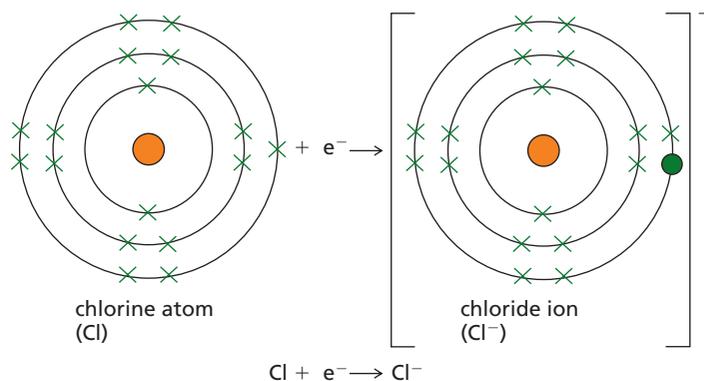
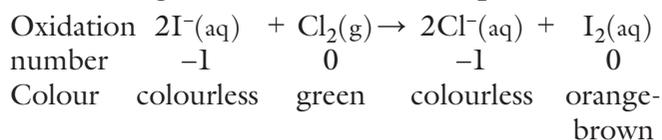


Figure 9.14 A chlorine atom gains an electron to form a chloride ion.

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 .



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

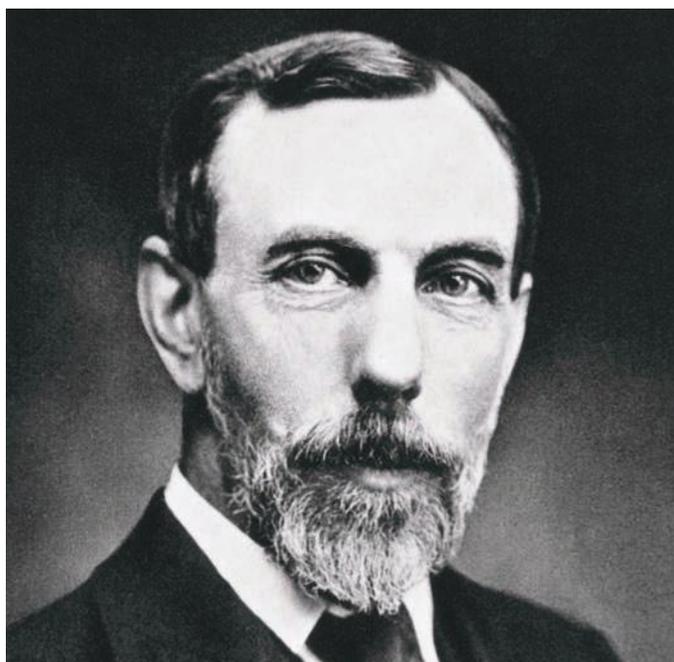
Helium, neon, argon, krypton, xenon and the radioactive element radon make up a most unusual group of non-metals, called the noble gases. They were all discovered after Mendeleev had published his periodic table. They were discovered between 1894 and 1900, mainly through the work of the British scientists 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

- Write word and balanced chemical equations for the reactions between:
 - bromine and potassium iodide solution
 - bromine and potassium chloride solution.
 If no reaction will take place, write 'no reaction' and explain why.
 - 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.
- 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.

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



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.

- 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, $\text{Cr}_2\text{O}_7^{2-}$, 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

- 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.
- Which groups in the Periodic Table contain:
 - only metals?
 - only non-metals?
 - 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: ${}_{17}^{35.5}\text{Cl}$, ${}_{35}^{80}\text{Br}$ and ${}_{53}^{127}\text{I}$.
- 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

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

- Write a word equation for the reaction of fluorine gas with sodium chloride solution.
 - Write a balanced chemical equation for the reaction, with state symbols.
 - 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.

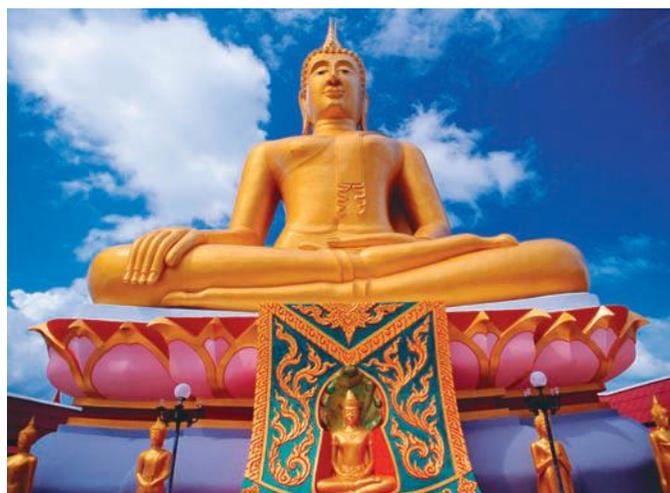


a Sodium burning in air/oxygen.

Figure 10.1



b Iron rusts when left unprotected.



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

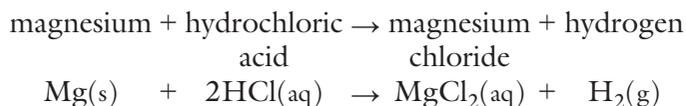
Sodium is soft and reacts violently with both air and water. Iron also reacts with air and water but much more slowly, forming rust. Gold, however, remains totally unchanged after many hundreds of years. Sodium is said to be more reactive than iron and, in turn, iron is said to be more reactive than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



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.

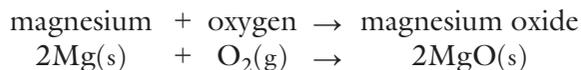
Reactivity series	Reaction with dilute acid	Reaction with air/oxygen	Reaction with water	Ease of extraction				
Potassium (K)	Produce H ₂ with decreasing vigour	Burn very brightly and vigorously	Produce H ₂ with decreasing vigour with cold water	Difficult to extract				
Sodium (Na)								
Calcium (Ca)	↓	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)								
[Hydrogen (H)]								
Copper (Cu)					Do not react with dilute acids	Do not react	Do not react with cold water or steam	Found as the element (native)
Silver (Ag)					↓	↓	↓	↓
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.

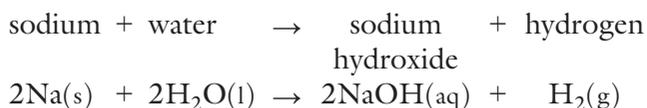
With air/oxygen

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



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.



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.

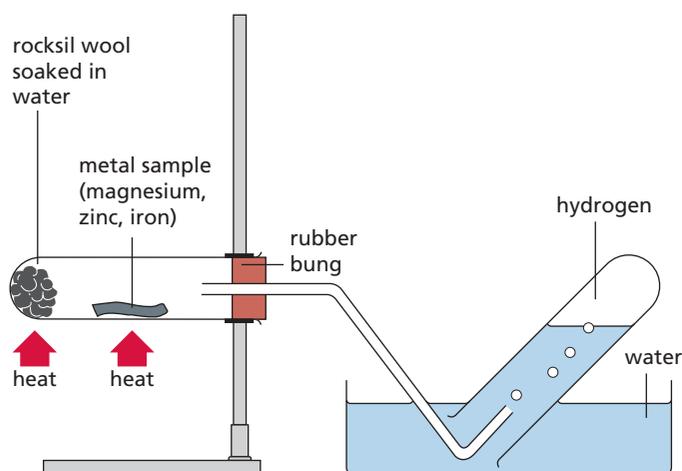
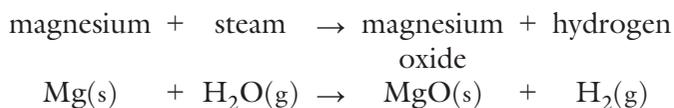


Figure 10.3 Apparatus used to investigate how metals such as magnesium react with steam.

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

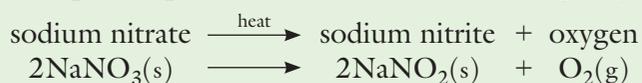
- Write balanced chemical equations for the reactions between:
 - iron and dilute hydrochloric acid
 - zinc and oxygen
 - calcium and water.
- Make a list of six things you have in your house made from copper or iron. Give a use for each of the other unreactive metals shown in the reactivity series.

Decomposition of metal nitrates, carbonates, oxides and hydroxides

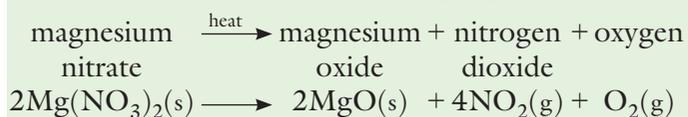
The ways in which metal nitrates, carbonates, oxides and hydroxides decompose can also be discussed in terms of the reactivity series of the metals. The decomposition processes are different, depending on the position of the metal in the reactivity series.

Metal nitrates

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



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



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:



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.

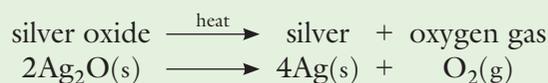


Table 10.2 Effect of heat on metal nitrates.

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 $2\text{KNO}_3(\text{s}) \longrightarrow 2\text{KNO}_2(\text{s}) + \text{O}_2(\text{g})$
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 $2\text{Ca}(\text{NO}_3)_2(\text{s}) \longrightarrow 2\text{CaO}(\text{s}) + 4\text{NO}_2(\text{g}) + \text{O}_2(\text{g})$ zinc nitrate → zinc oxide + nitrogen(IV) oxide + oxygen $2\text{Zn}(\text{NO}_3)_2(\text{s}) \longrightarrow 2\text{ZnO}(\text{s}) + 4\text{NO}_2(\text{g}) + \text{O}_2(\text{g})$ copper(II) nitrate → copper oxide + nitrogen(IV) oxide + oxygen $2\text{Cu}(\text{NO}_3)_2(\text{s}) \longrightarrow 2\text{CuO}(\text{s}) + 4\text{NO}_2(\text{g}) + \text{O}_2(\text{g})$
Silver Platinum Gold	Ag Pt Au	Decompose to the metal releasing oxygen and nitrogen(IV) oxide	silver nitrate → silver + nitrogen(IV) oxide + oxygen $2\text{AgNO}_3(\text{s}) \longrightarrow 2\text{Ag}(\text{s}) + 2\text{NO}_2(\text{g}) + \text{O}_2(\text{g})$

Table 10.3 Effect of heat on metal carbonates.

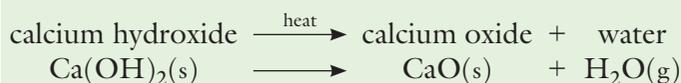
Metal	Symbol	Effect of heat on carbonate	Equation (some examples)
Potassium Sodium	K Na	No reaction	
Calcium Magnesium Aluminium Zinc Lead Copper	Ca Mg Al Zn Pb Cu	Form the oxide and release carbon dioxide with increasing ease	calcium carbonate → calcium oxide + carbon dioxide $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$ zinc carbonate → zinc oxide + carbon dioxide $\text{ZnCO}_3(\text{s}) \rightarrow \text{ZnO}(\text{s}) + \text{CO}_2(\text{g})$ copper(II) carbonate → copper(II) oxide + carbon dioxide $\text{CuCO}_3(\text{s}) \rightarrow \text{CuO}(\text{s}) + \text{CO}_2(\text{g})$
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 $\text{Ca}(\text{OH})_2(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{H}_2\text{O}(\text{g})$ zinc hydroxide → zinc oxide + water $\text{Zn}(\text{OH})_2(\text{s}) \rightarrow \text{ZnO}(\text{s}) + \text{H}_2\text{O}(\text{g})$ copper hydroxide → copper oxide + water $\text{Cu}(\text{OH})_2(\text{s}) \rightarrow \text{CuO}(\text{s}) + \text{H}_2\text{O}(\text{g})$
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.

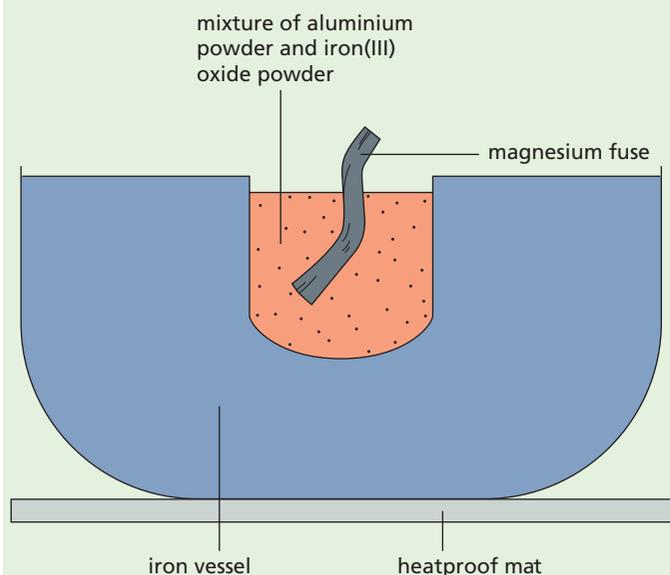
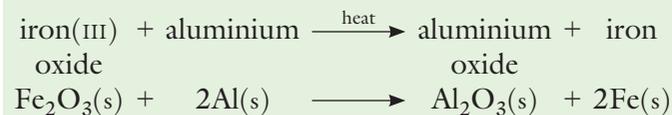
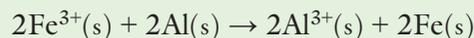


Figure 10.6 Themit reaction mixture apparatus.

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.



or



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 **Themit 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.

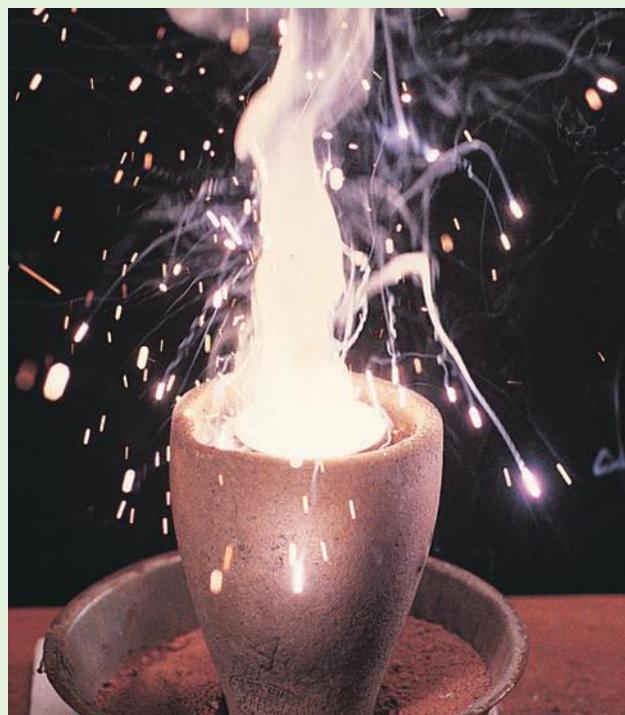


Figure 10.7 The Themit reaction in a laboratory. The same reaction is used to weld damaged railway lines.

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.

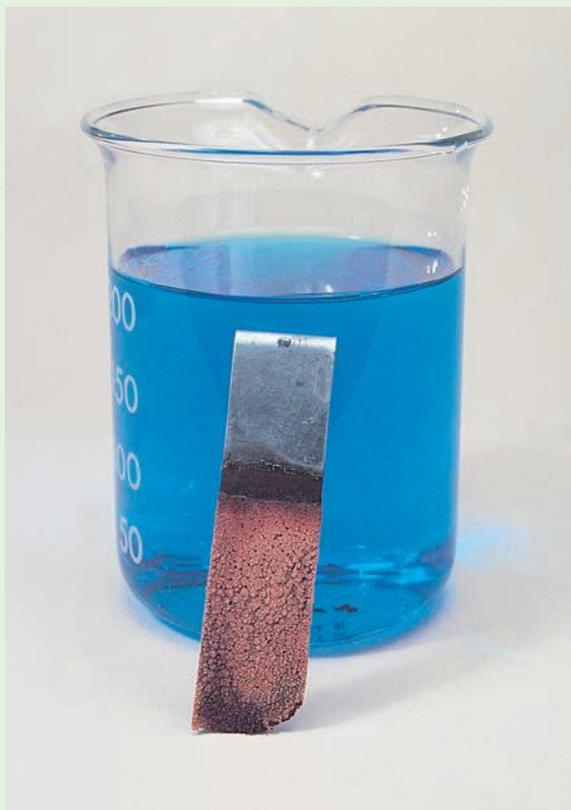
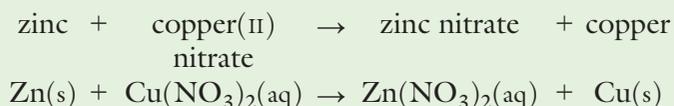
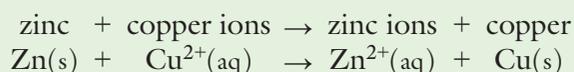


Figure 10.8 Zinc displaces copper.

The ionic equation for this reaction is:



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

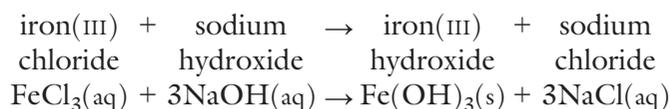
✗ = no reaction

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

Identifying metal ions

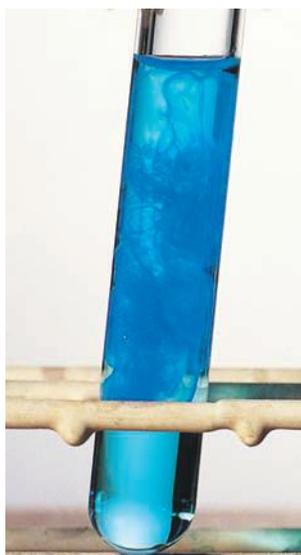
When an alkali dissolves in water, it produces hydroxide ions. It is known that most metal hydroxides are insoluble. So if hydroxide ions from a solution of an alkali are added to a solution of a metal salt, an insoluble, often coloured, metal hydroxide is precipitated from solution (Figure 10.9).

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





a Iron(III) hydroxide is precipitated.



b Copper(II) hydroxide is precipitated.

Figure 10.9

The ionic equation for this reaction is:

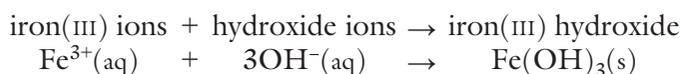


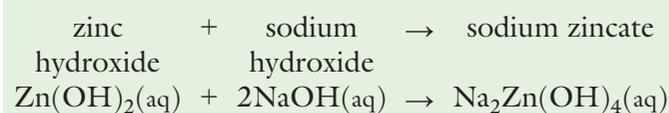
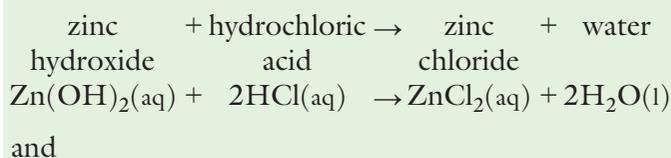
Table 10.5 shows the effects of adding a few drops of sodium hydroxide solution to solutions containing various metal ions, and of adding an excess. The colours of the insoluble metal hydroxides can be used to identify the metal cations present in solution. In some cases the precipitate dissolves in 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

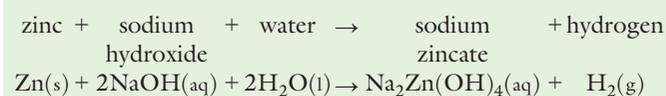
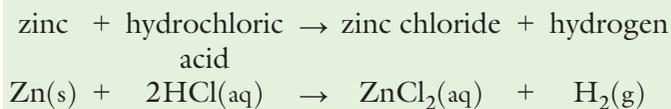
Amphoteric hydroxides and oxides

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,

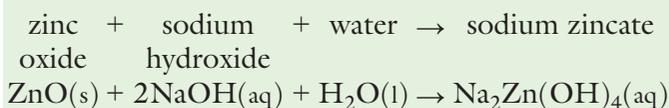
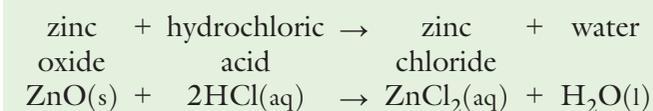


Other amphoteric hydroxides are lead hydroxide ($\text{Pb}(\text{OH})_2$) and aluminium hydroxide ($\text{Al}(\text{OH})_3$). We can use this sort of behaviour to help identify metal cations, as their hydroxides are soluble in strong bases.

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



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.



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

- Write equations for the reaction between:
 - aluminium and moderately concentrated hydrochloric acid
 - aluminium and moderately concentrated sodium hydroxide (producing sodium aluminate, $\text{NaAl}(\text{OH})_4$).
- Write ionic equations for the reactions which take place to produce the metal hydroxides shown in Table 10.5.
- 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	$\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$	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.

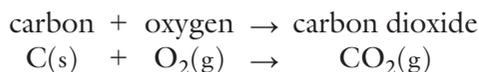
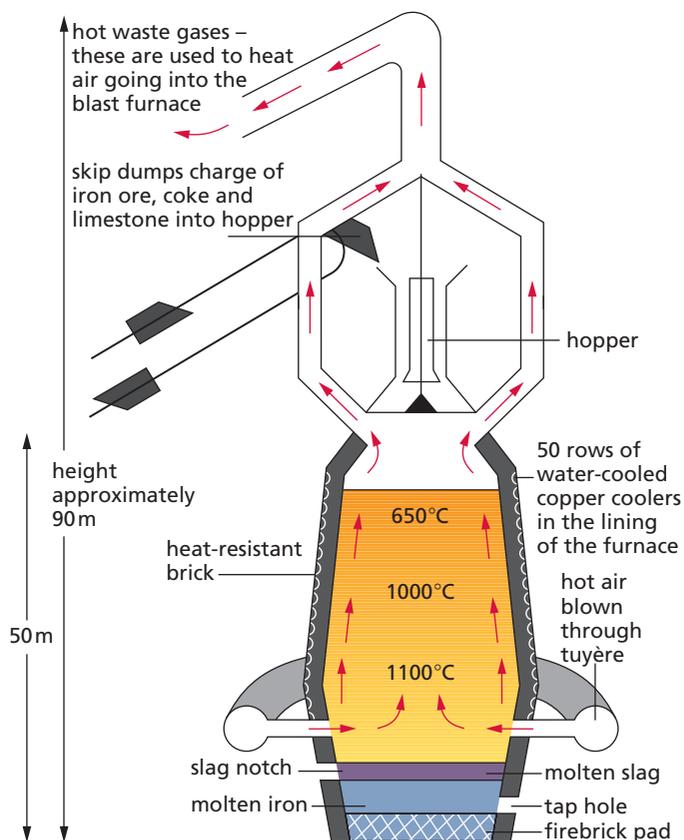
Extraction of fairly reactive metals

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

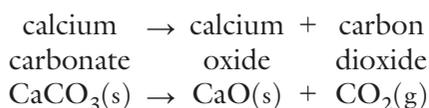
Iron is extracted mainly from its oxides, haematite (Fe_2O_3) and magnetite (Fe_3O_4), in a blast furnace (Figures 10.12 and 10.13). These ores contain at least 60% iron. The iron ores used are a blend of those extracted in Australia, Canada, Sweden, Venezuela and Brazil. The blast furnace is a steel 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.

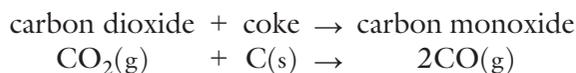
**Figure 10.12** A blast furnace.**Figure 10.13** Cross-section of a blast furnace.

A number of chemical reactions then follow.

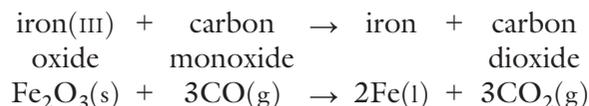
- The limestone begins to decompose:



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

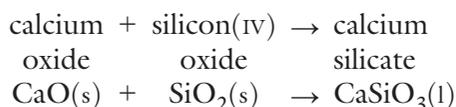


- Carbon monoxide is a reducing agent (Chapter 2, p. 14). It rises up the furnace and reduces the iron(III) oxide ore. This takes place at a temperature of around 700 °C:



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.



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 (SiO₂), 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).

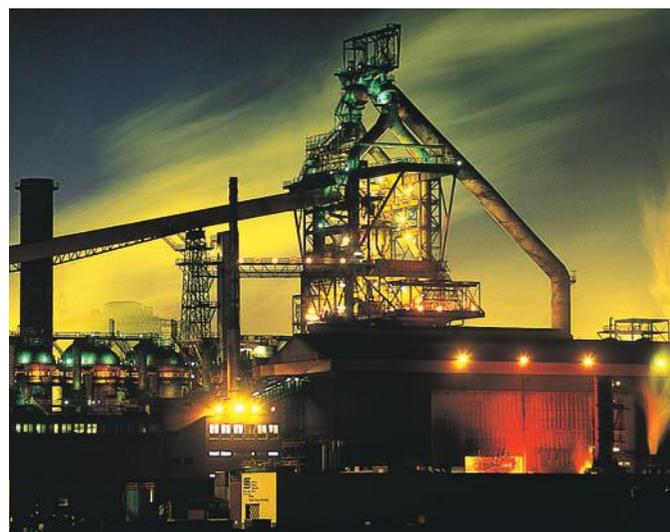


Figure 10.15 Redcar steelworks at Teesside, UK.

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:

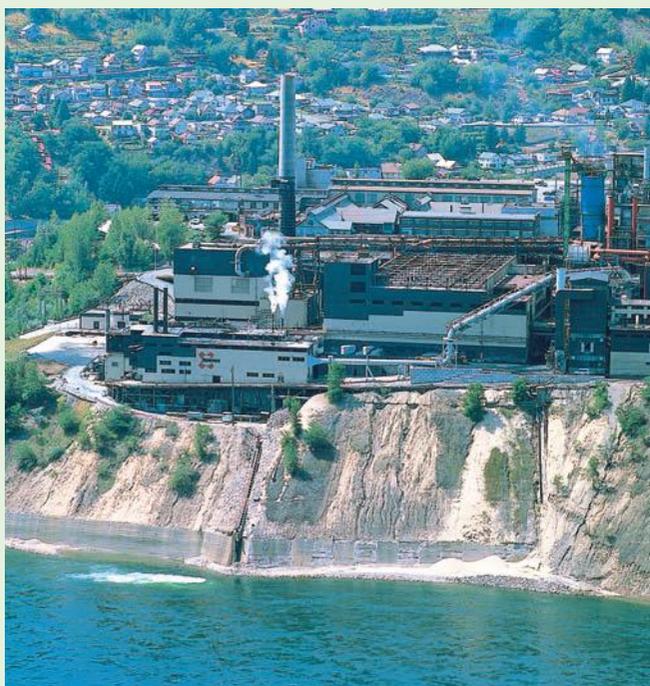
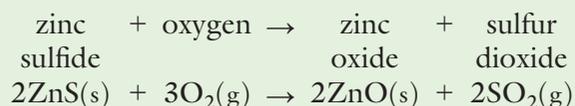
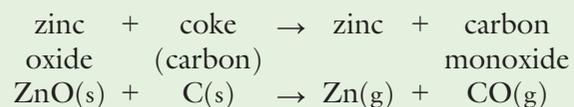


Figure 10.16 Zinc is extracted in a furnace like this one.

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:



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 ($\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$). 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,

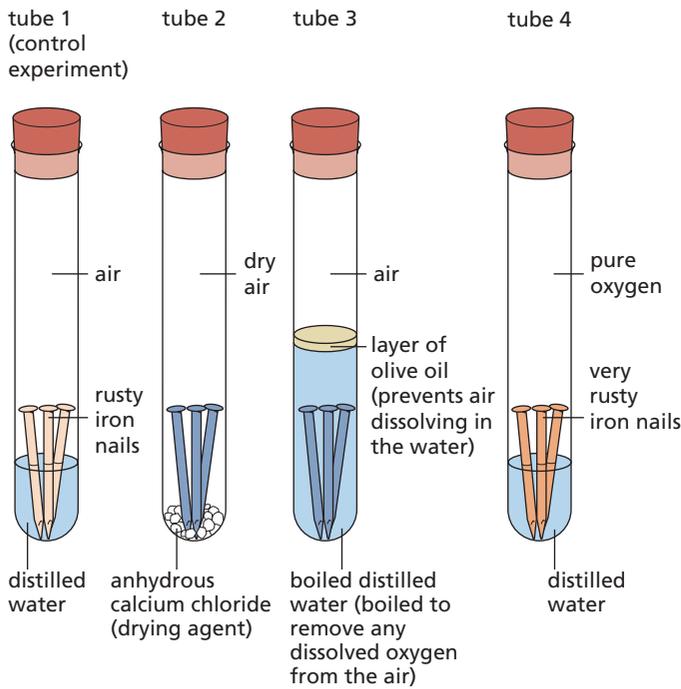


Figure 10.20 Rusting experiment with nails.



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



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.

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

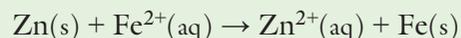
Some steel girders, used in the construction of bridges and buildings, are **galvanised**. Coal bunkers and steel dustbins are also galvanised. This involves dipping the object into molten zinc (Figure 10.24). The thin layer of the more reactive zinc metal coating the steel object slowly corrodes and 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:

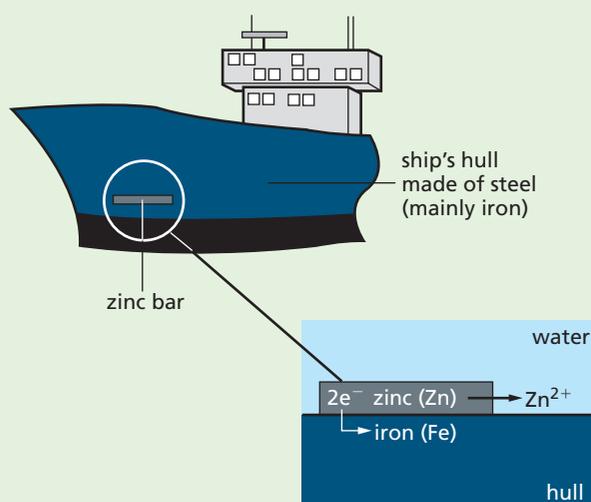


As long as some of the zinc bars remain in contact with the iron structure, the structure will be protected from rusting. When the zinc runs out,

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 ($\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2$).

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.

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.

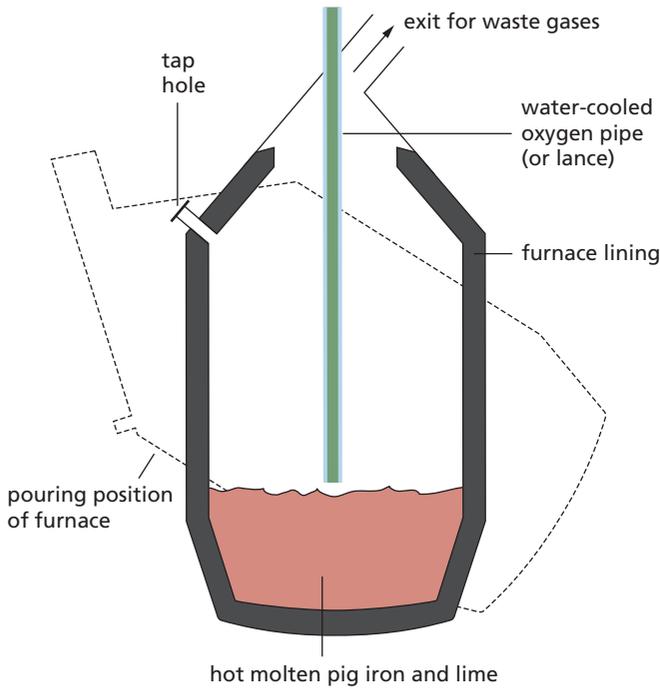


Figure 10.27 A stainless steel exhaust system. Why do you think more people are buying these exhaust systems?

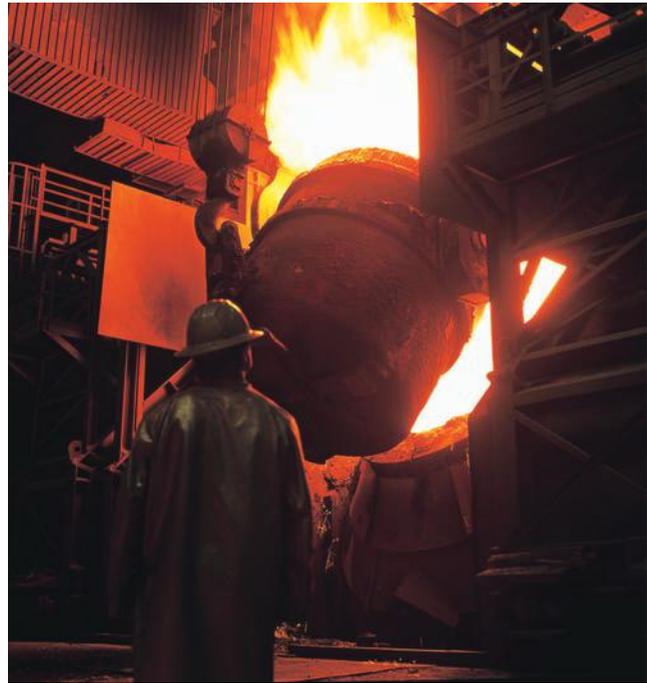
- Carbon is oxidised to carbon monoxide and carbon dioxide, while sulfur is oxidised to sulfur dioxide. These escape as gases.
- Silicon and phosphorus are oxidised to silicon(IV) oxide and phosphorus pentoxide, which are solid oxides.
- Some calcium oxide (lime) is added to remove these solid oxides 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

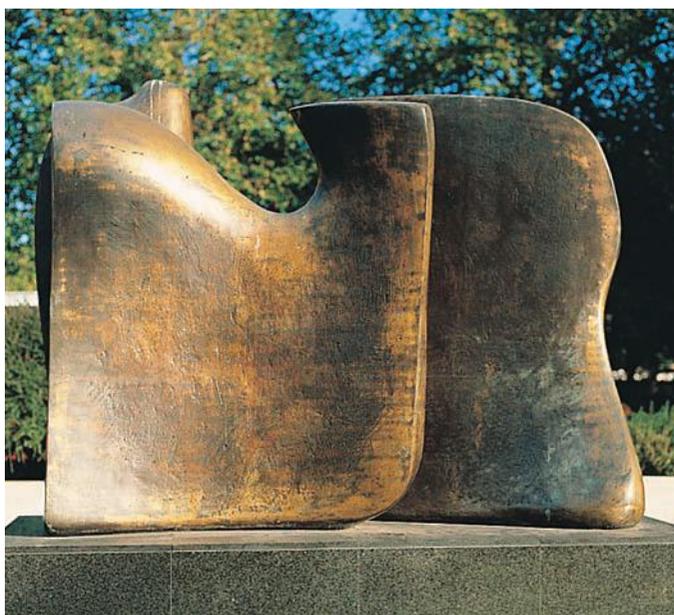
Steel recycling

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.

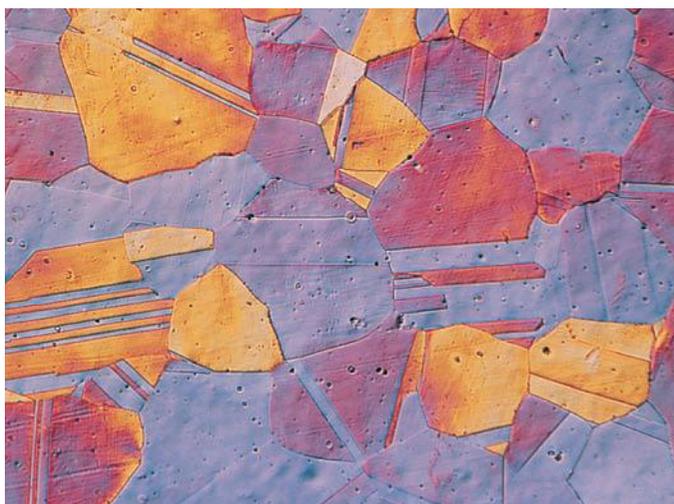
Alloys to order

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).

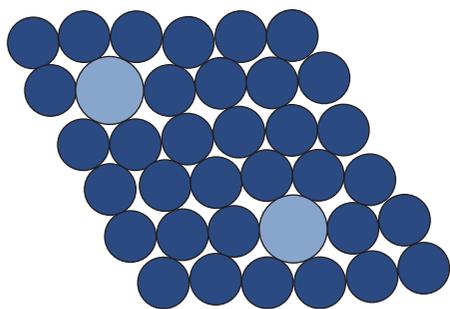
Table 10.8 shows some of the more common alloys, together with some of their uses.



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.

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.

Figure 10.29

Checklist

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

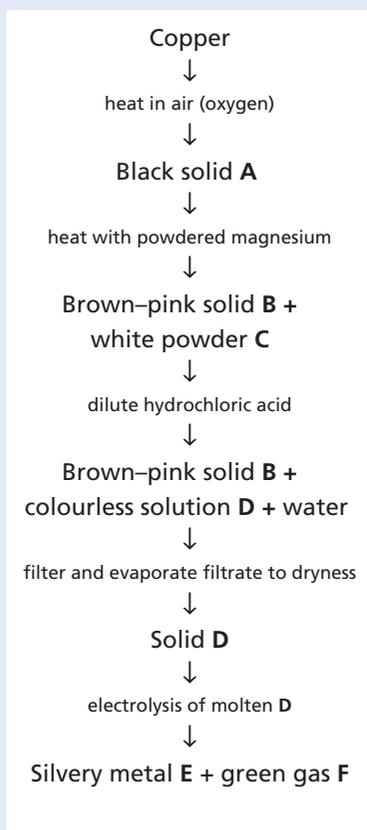
- **Alloy** Generally, a mixture of two or more metals (for example, brass is an alloy of zinc and copper) or of a metal and a non-metal (for example, steel is an alloy of iron and carbon, sometimes with other metals included). They are formed by mixing the molten substances thoroughly. Generally, it 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

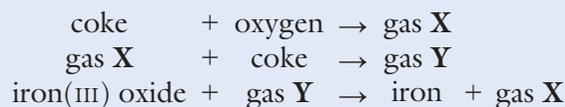
- 1 Use the following list of metals to answer the questions a to i: iron, calcium, potassium, gold, aluminium, magnesium, sodium, zinc, platinum, titanium.
- Which of the metals is found native?
 - Which of the metals is found in nature as the ore:
 - rock salt?
 - rutile?
 - Which metal has a carbonate found in nature called chalk?
 - Which of the metals will not react with oxygen to form an oxide?
 - Which of the metals will react violently with cold water?
 - Choose one of the metals in your answer to e and write a balanced chemical equation for the reaction which takes place.
 - Which of the metals has a protective oxide coating on its surface?
 - Which of the metals reacts very slowly with cold water but extremely vigorously with steam?
 - Which of the metals is used to galvanise iron?

2

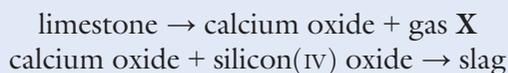


- a Name and give the formulae of the substances A to F.

- Write balanced chemical equations for the reactions in which:
 - black solid A was formed
 - white powder C and brown-pink solid B were formed
 - colourless solution D was formed.
 - The reaction between black solid A and magnesium is a redox reaction. With reference to this reaction, explain what you understand by this statement.
 - Write anode and cathode reactions for the processes which take place during the electrolysis of molten D.
 - Suggest a use for:
 - brown-pink solid B
 - silvery metal E
 - green gas F.
- 3 Explain the following:
- metals such as gold and silver occur native in the Earth's crust
 - the parts of shipwrecks made of iron rust more slowly in deep water
 - zinc bars are attached to the structure of oil rigs to prevent them from rusting
 - copper roofs quickly become covered with a green coating when exposed to the atmosphere
 - 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.



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



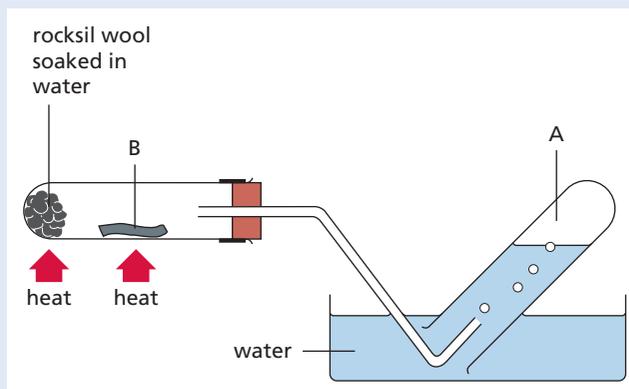
- Give the chemical names for limestone and slag.
- Write balanced chemical equations for the reactions shown above.

- (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.
- Explain the meaning of the term alloy as applied to steel.
 - Name two impurities, other than carbon, which are present in cast iron and which are removed completely during the steel manufacture.
 - Describe the method of steel manufacture used which removes the impurities referred to in **b**.
 - Name two metallic elements which may be added to the basic oxygen furnace to produce different varieties of steel.
 - Give two uses of stainless steel.
 - Give two advantages of stainless steel compared with cast iron.
- 6 The metal titanium is extracted from its ore rutile (TiO_2).
- Describe the processes involved in the extraction of titanium from rutile.
 - Name the area of the Periodic Table in which titanium is found.
 - The final part of the process involves the following reaction:

$$\text{TiCl}_4(l) + 4\text{Na}(l) \rightarrow 4\text{NaCl}(s) + \text{Ti}(s)$$
 - Calculate the quantity of titanium produced from 19 tonnes of titanium(IV) chloride (TiCl_4).
 - How much sodium would have to be used to extract the titanium from the 19 tonnes of TiCl_4 used in (i)?
 - 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.



- Name and give the formula of gas **A**.
 - Name the product formed when gas **A** burns in air.
 - Write a balanced chemical equation for this reaction.
 - Name white solid **B**.
 - Write a balanced chemical equation to represent the reaction between magnesium and steam.
 - Name two other metals which could be safely used to replace zinc and produce another sample of gas **A**.
 - 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.
- $$\text{Cu}_2\text{S}(s) + \text{O}_2(g) \rightarrow 2\text{Cu}(s) + \text{SO}_2(g)$$
- What is the name of the chemical process by which copper(I) sulfide is converted to copper?
 - At what temperature does this process take place?
 - 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?
 - Calculate the quantity of copper(I) sulfide required to produce 1270 tonnes of copper. (A_r : O = 16; S = 32, Cu = 63.5)
 - 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.

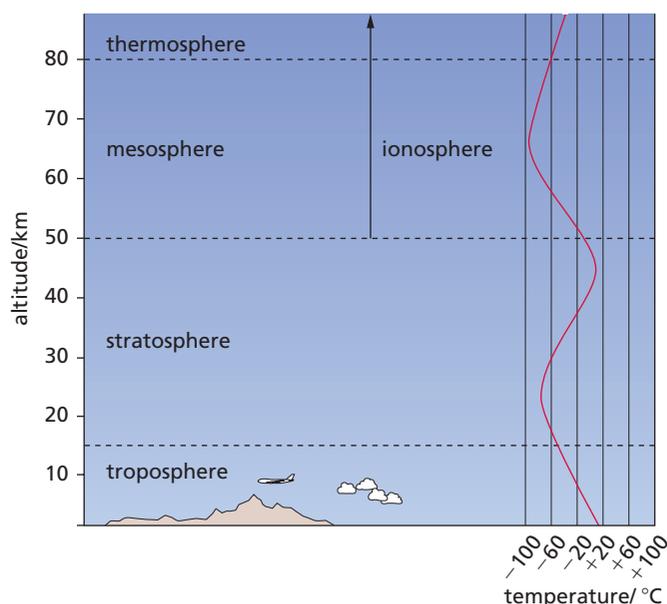


Figure 11.2 The Earth's atmosphere.

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:

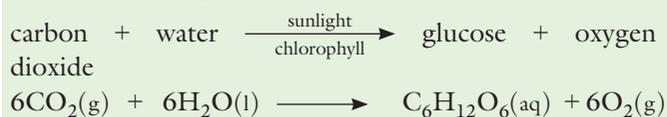


Figure 11.4 This algae is related to that early organism produced all those years ago!

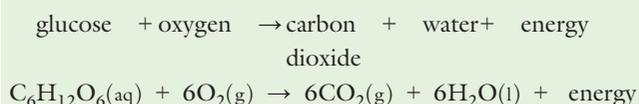
These organisms used the light from the Sun to produce their own food, and oxygen was released into the atmosphere as a waste product of **photosynthesis**. This process also acted to reduce the amount of CO_2 in the atmosphere.

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, $\text{O}_3(\text{g})$. 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:



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.

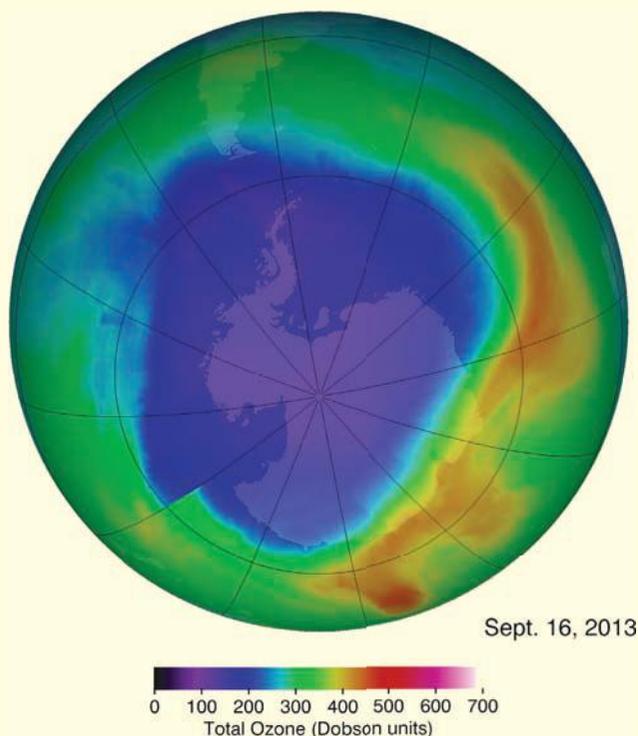


Figure 11.5 This diagram comes from NASA's ozone monitoring programme TOMS (Total Ozone Mapping Spectrometer). The ozone hole over the Antarctic (shown in purple and pink on the diagram) is largest in the Antarctic spring. Note: Dobson Units are a measure of the total amount of ozone in a vertical column 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:

$$\frac{21}{100} \times 100 = 21\%$$

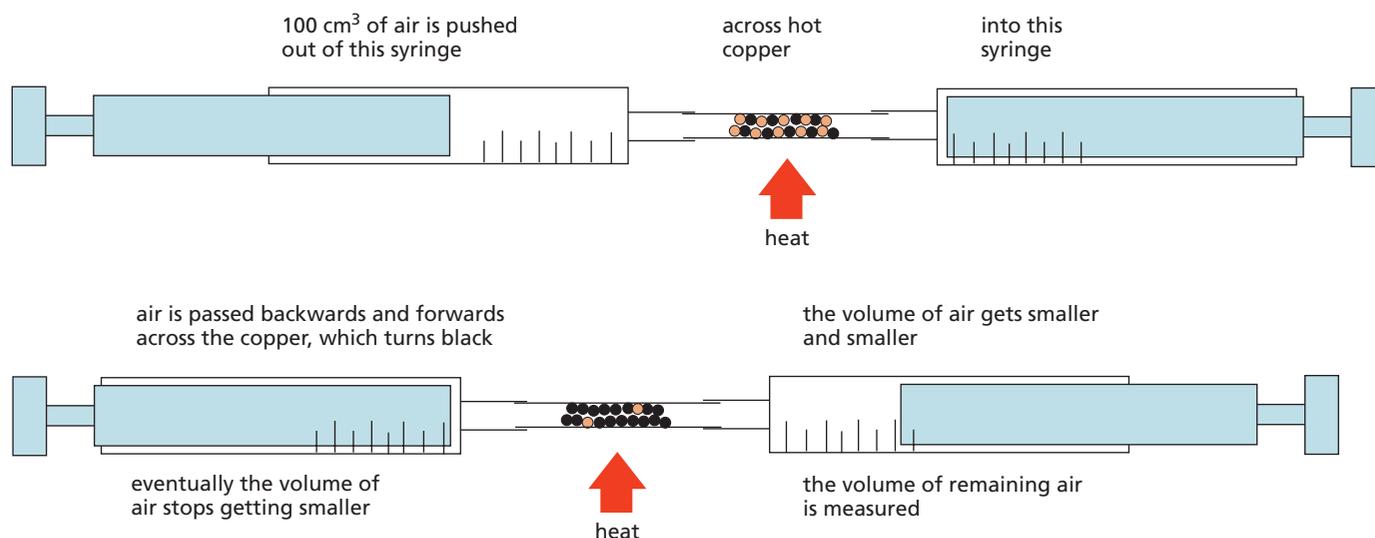


Figure 11.6 This apparatus can be used to find out the volume of oxygen gas in the air.



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).

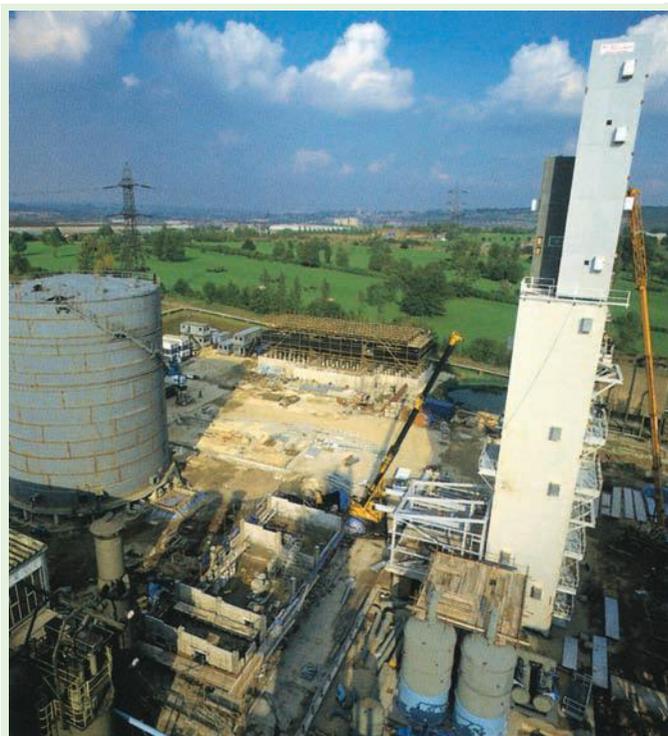


Figure 11.8 Large amounts of gases are obtained from the fractional distillation of liquid air in industrial plants such as this.

- 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/ $^{\circ}\text{C}$
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	<ul style="list-style-type: none"> • 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	<ul style="list-style-type: none"> • 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	<ul style="list-style-type: none"> 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	<ul style="list-style-type: none"> 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	<ul style="list-style-type: none"> 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	<ul style="list-style-type: none"> 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.



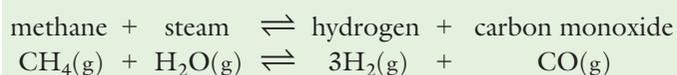
Figure 11.13 Fritz Haber (1868–1934).

Obtaining nitrogen

The nitrogen needed in the Haber process is obtained from the atmosphere by fractional distillation of liquid air (p. 174).

Obtaining hydrogen

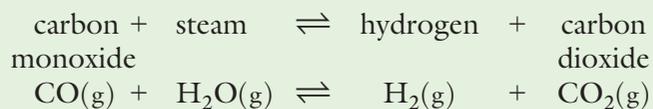
The hydrogen needed in the Haber process is obtained from the reaction between methane and steam.



This process is known as **steam re-forming**.

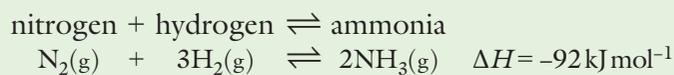
This reaction is a **reversible** reaction and special conditions are employed to ensure that the reaction proceeds to the right (the forward reaction), producing hydrogen and carbon monoxide. The process is carried out at a temperature of 750 °C, at a pressure of 30 atmospheres with a catalyst of nickel. 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.



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:



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

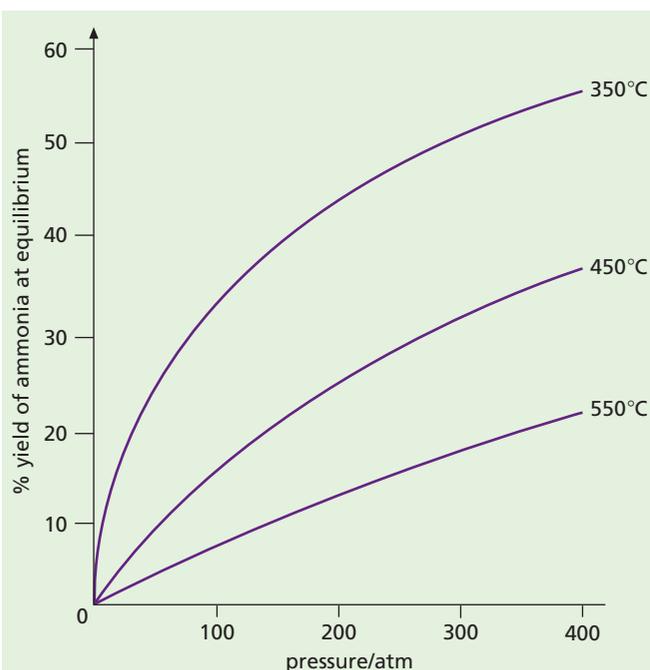


Figure 11.14 Yields from the Haber process.

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.

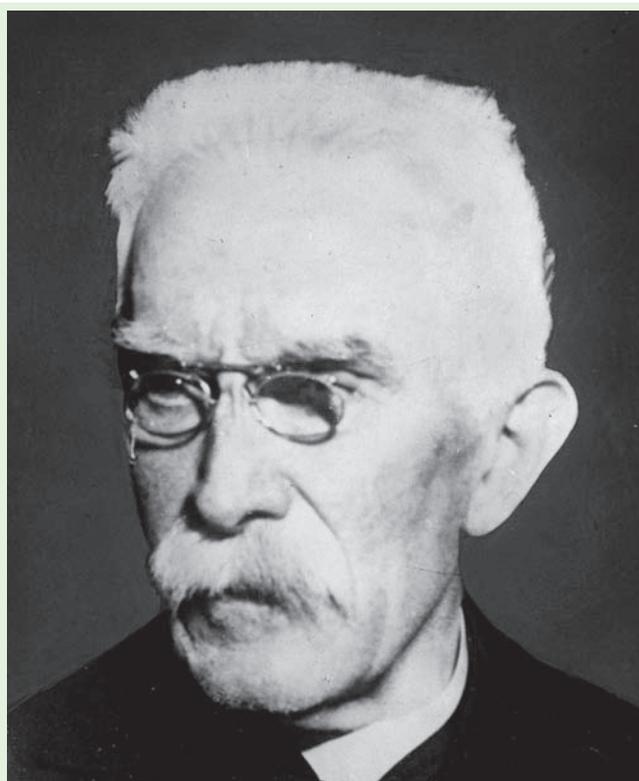


Figure 11.15 Henri Le Chatelier (1850–1936).

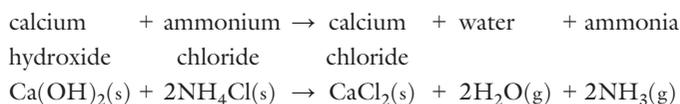
Worldwide, in excess of 140 million tonnes of ammonia are produced by the Haber process each year.

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.



Water vapour is removed from the ammonia gas by passing the gas formed through a drying tower containing calcium oxide (Figure 11.16).

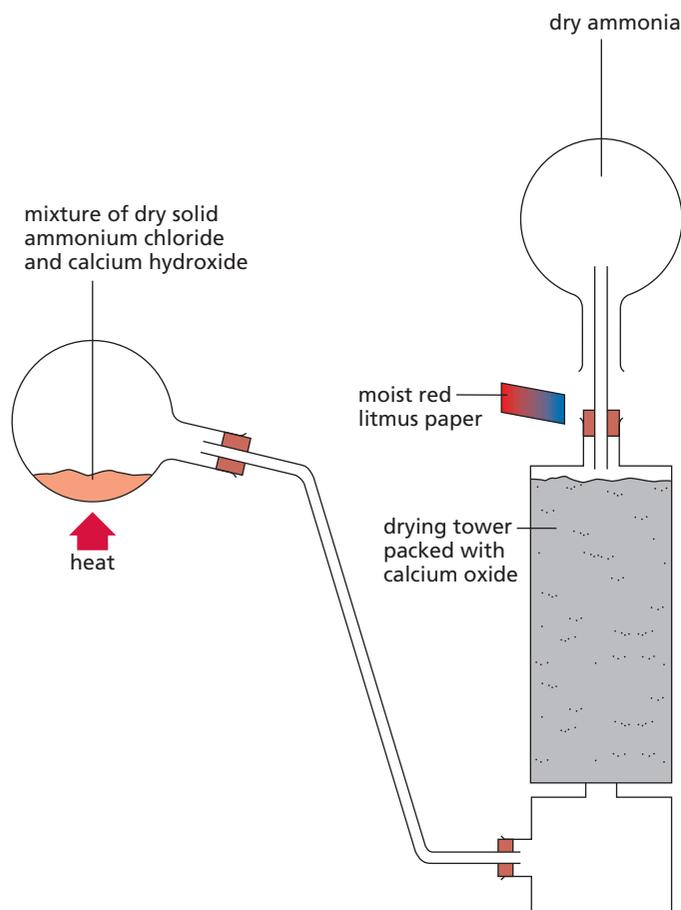


Figure 11.16 Laboratory production of ammonia gas.

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).

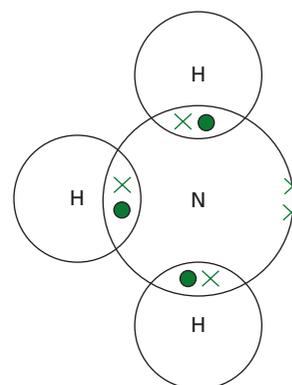
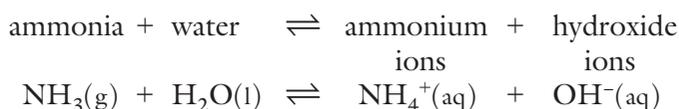


Figure 11.17 The ammonia molecule.

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.



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).

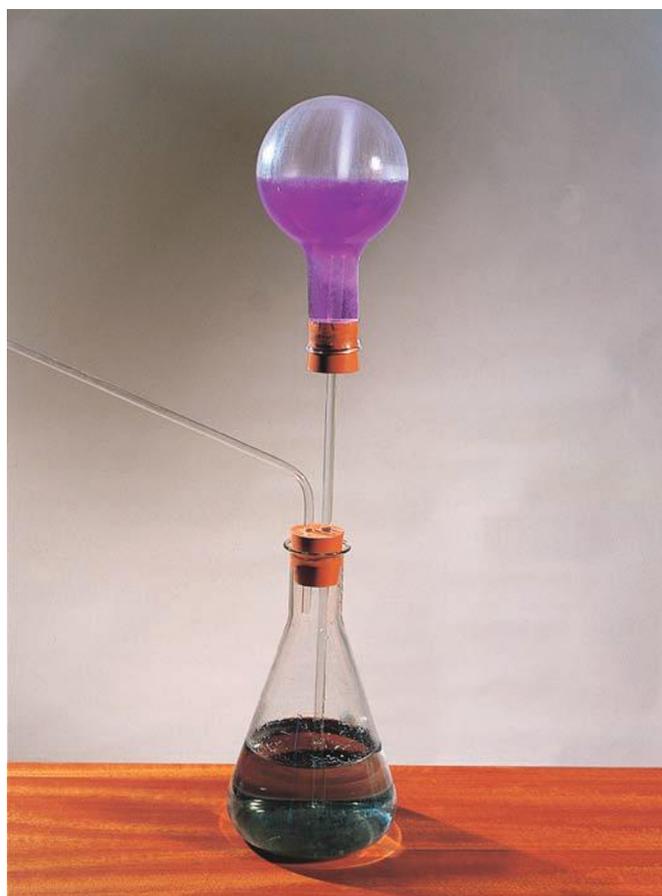
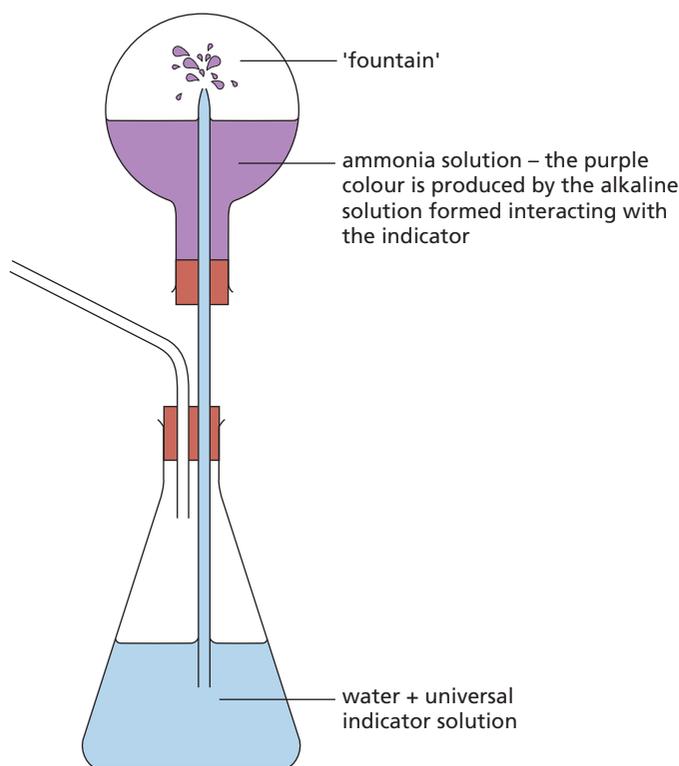


Figure 11.18 The fountain flask experiment.

Table 11.6 Identifying metal ions using aqueous ammonia.

Metal ion	With a few drops of ammonia solution	With excess ammonia solution
$\text{Cu}^{2+}(\text{aq})$	Gelatinous blue precipitate	Precipitate dissolves to give a deep blue solution
$\text{Fe}^{2+}(\text{aq})$	Dirty green precipitate	Dirty green precipitate remains
$\text{Fe}^{3+}(\text{aq})$	Rust brown precipitate	Rust brown precipitate remains
$\text{Al}^{3+}(\text{aq})$	White precipitate	White precipitate remains
$\text{Zn}^{2+}(\text{aq})$	White precipitate	White precipitate dissolves to give a colourless solution
$\text{Ca}^{2+}(\text{aq})$	Faint white precipitate	Faint white precipitate remains
$\text{Cr}^{3+}(\text{aq})$	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 $\text{Al}^{3+}(\text{aq})$, $\text{Ca}^{2+}(\text{aq})$ and $\text{Cu}^{2+}(\text{aq})$.

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.

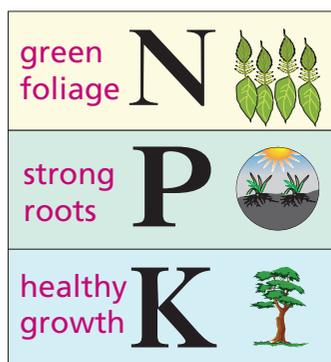
Fertiliser	Formula
Ammonium nitrate	NH_4NO_3
Ammonium phosphate	$(\text{NH}_4)_3\text{PO}_4$
Ammonium sulfate	$(\text{NH}_4)_2\text{SO}_4$
Urea	$\text{CO}(\text{NH}_2)_2$

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 ($(\text{NH}_4)_3\text{PO}_4$) 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.

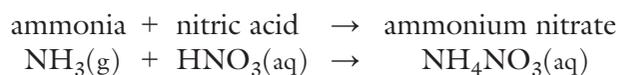
Figure 11.19



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.



The nitrogen cycle

The vital importance of nitrogen to both plants and animals can be summarised by the **nitrogen cycle** (Figure 11.21).

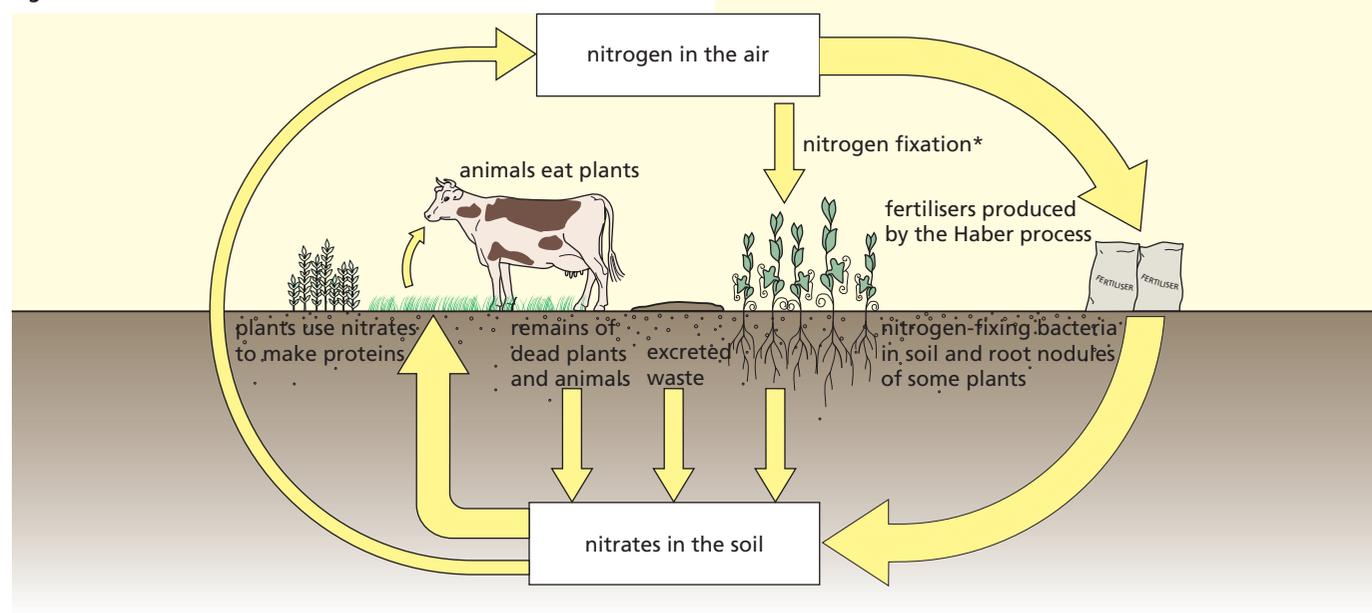


Figure 11.21 The nitrogen cycle.

***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

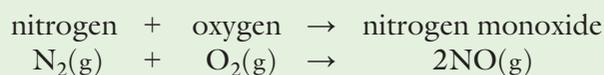
The two major resources considered in this chapter, water and air, are essential to our way of life and our very existence. Water and air make up the environment of a living organism. The environment is everything in the surroundings of an organism that could possibly influence it. Humans continually 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:

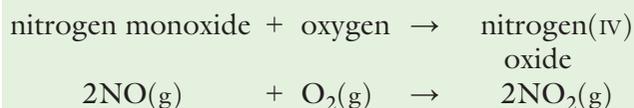
- carbon monoxide – toxic gas
- sulfur dioxide – major irritant and constituent of acid rain
- hydrocarbons – carcinogenic
- oxides of nitrogen – major irritant, acid rain and photochemical smog
- lead compounds – toxic.

Nitrogen monoxide is formed by the reaction of nitrogen and oxygen inside the internal combustion engine at high temperatures.



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).



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



Figure 11.23 Particulates produced by diesel engines are a real health problem worldwide.

have been associated with a variety of lung complaints in both adults and children.)

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.

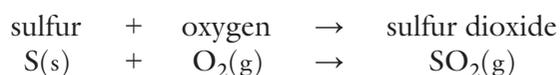
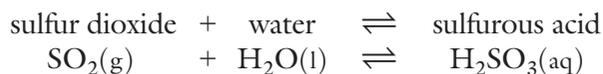


Figure 11.24 Sulfur dioxide is a major pollutant produced by industry.

This sulfur dioxide gas dissolves in rainwater to form the weak acid, sulfurous acid (H_2SO_3).



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).

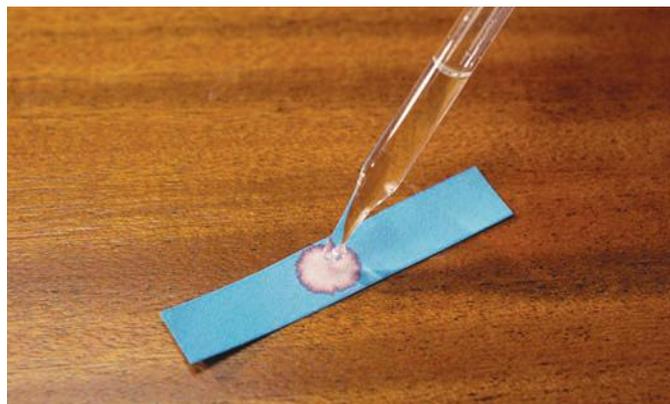


Figure 11.27 Liquid water turns to steam at 100°C and to ice at 0°C .

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).

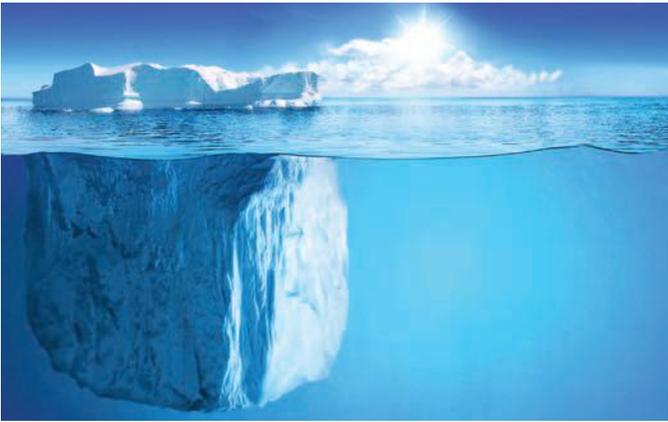


Figure 11.29 When water freezes, its density falls; this is why icebergs float.

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.

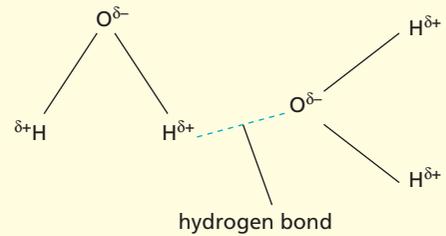


Figure 11.30 The unusual properties of water can be explained by hydrogen bonding.

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.

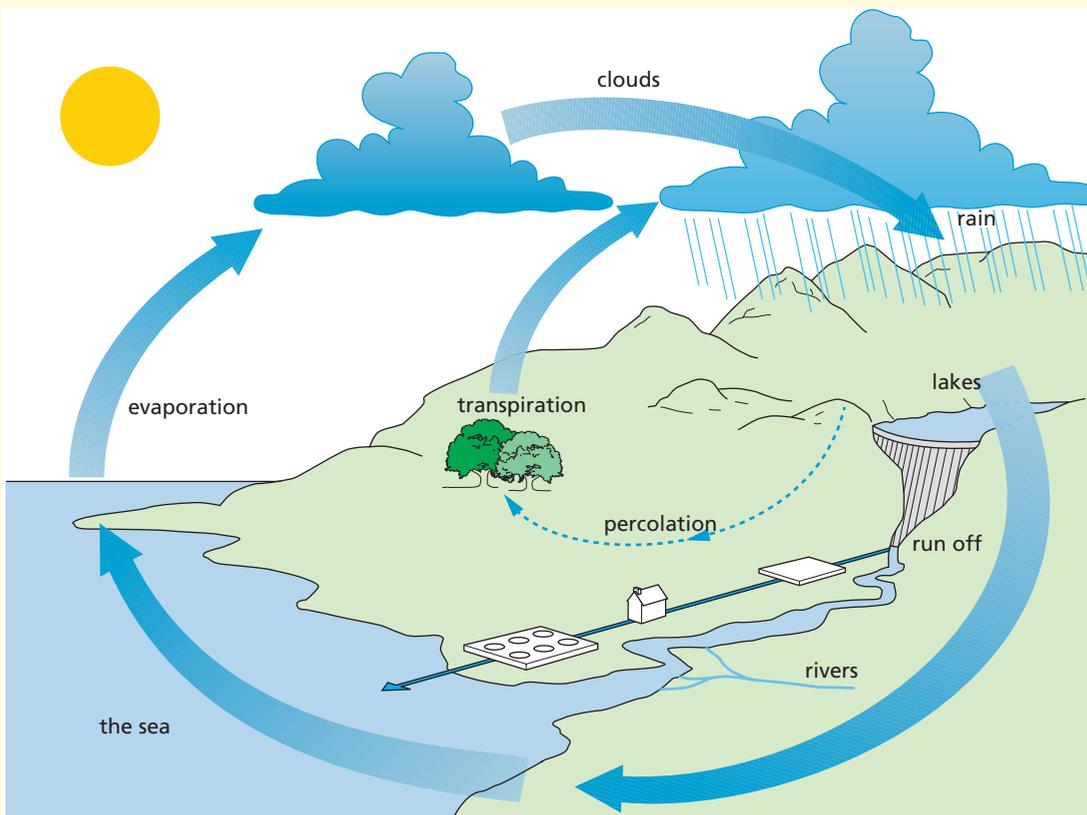


Figure 11.31 The water cycle.

- 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.

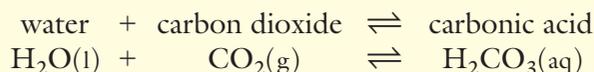
Some water is diverted from rivers into reservoirs and purified (p. 190). This water is then used in industry and the home. After use it is dirty and has to be treated at the sewage works before being returned to a river or the sea. Eventually, all the water finds its way into the sea. In this way all the water that left the sea eventually returns there, and the cycle can continue.

Question

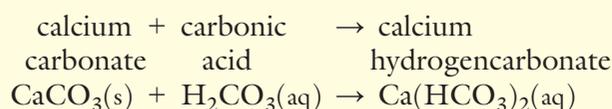
- 1 Construct a simplified version of the water cycle using 'key words' in boxes and the 'processes involved' over linking arrows.

● 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).



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.



Some of the rock strata may contain gypsum (calcium sulfate, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$), anhydrite (CaSO_4)

or kieserite ($\text{MgSO}_4 \cdot \text{H}_2\text{O}$), 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).

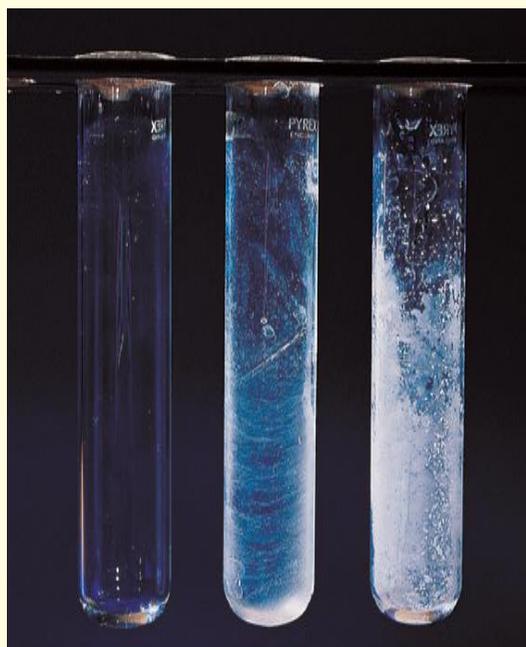
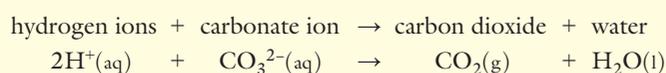


Figure 11.32 Evaporating pure water leaves no deposit, while temporary or permanent hard water leaves white deposits behind.

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:



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).

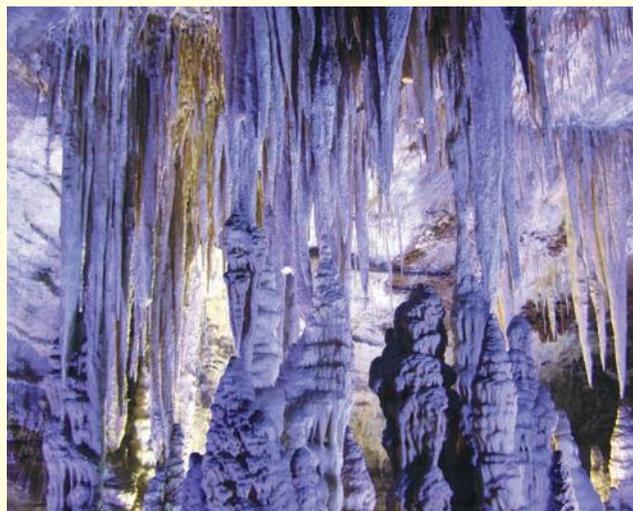
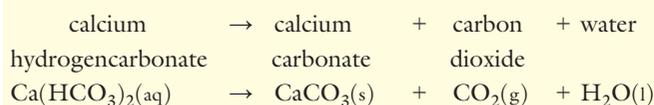


Figure 11.34 Stalactites and stalagmites have formed over hundreds of thousands of years.

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).

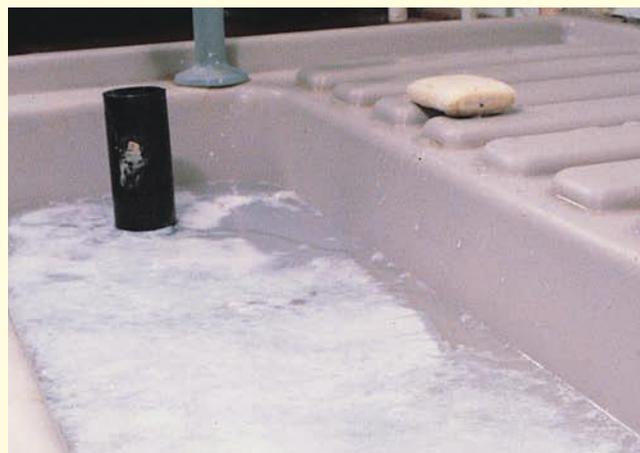
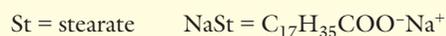
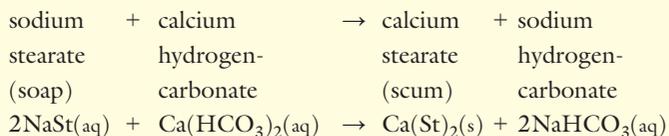


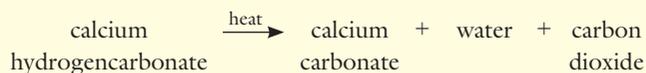
Figure 11.35 Soap and hard water form scum.

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.



● 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).

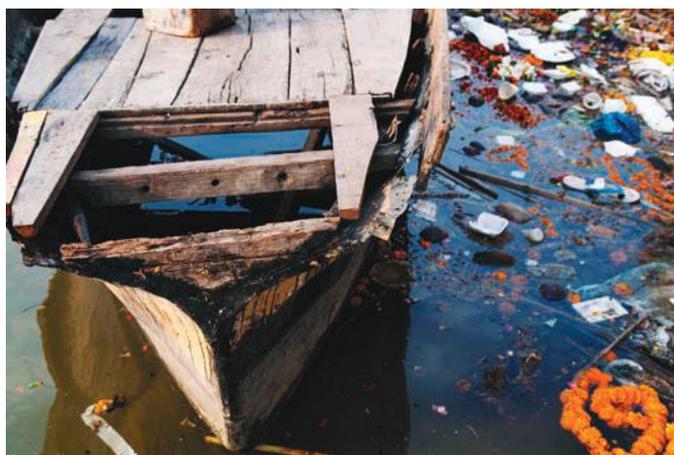


Figure 11.37 A badly polluted river.

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.

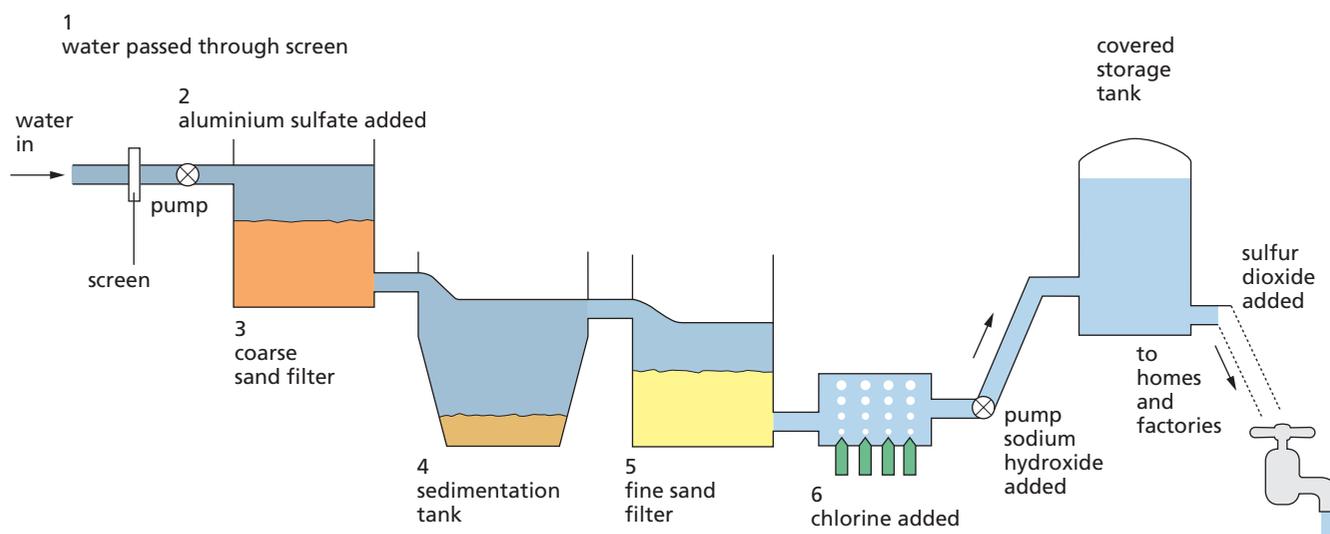


Figure 11.39 The processes involved in water treatment.

- 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.

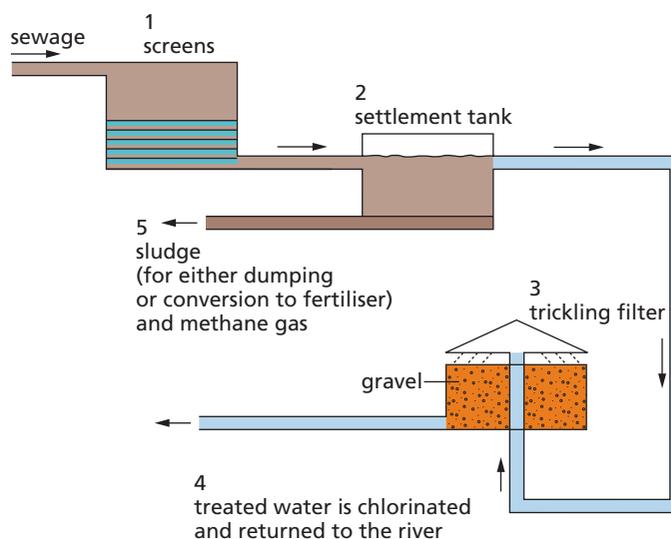


Figure 11.41 The processes involved in sewage treatment.

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	c sodium hydroxide
b chlorine	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) $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$	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) $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g})$ $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g})$	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.

- Why was the apparatus allowed to cool back to room temperature before the final volume reading was taken?
 - Using the information given above, calculate the percentage volume reduction which has taken place.
 - Explain briefly why there is a change in volume.
 - What observation given above supports your explanation in c? Write a balanced chemical equation for any reaction which has occurred.
 - Give the name of the main residual gas at the end of the experiment.
 - Would you expect the copper to have increased or decreased in mass during the experiment? Explain your answer.
- 2
- Oxygen has an atomic number of 8 and mass number 16. What is the electron structure of the oxygen atom?
 - How many electrons, neutrons and protons are there in the oxygen atom?
 - Oxygen molecules are diatomic. Explain the meaning of this term.
 - 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.
- Air is a mixture of elements and compounds.
 - The percentage of carbon dioxide in the atmosphere does not significantly vary from 0.04%.
 - 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:

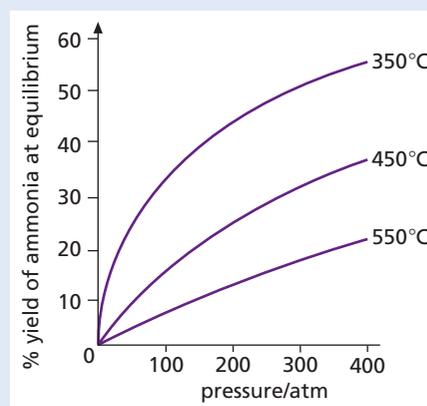
- in chemical _____, for example in the production of ethanol from ethene
- as a _____ to dissolve things
- as a _____ in some chemical processes.

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
- Explain what is meant by the term 'pollution' with reference to air and water.
 - Name an air pollutant produced by the burning of coal.
 - Name a different air pollutant produced by the combustion of petrol in a car engine.
 - Some of our drinking water is obtained by purifying river water.
 - Would distillation or filtration produce the purest water from river water? Give a reason for your answer.
 - 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).
 - Power stations produce warm water. This causes thermal pollution as this warm water is pumped into nearby rivers.
 - Why do power stations produce such large quantities of warm water?
 - What effect does this warm water have on aquatic life?

Fertiliser	Formula	% nitrogen
Ammonia solution	NH ₃	82.4
Calcium nitrate	Ca(NO ₃) ₂	17.1
Nitram	NH ₄ NO ₃	35.0
Sodium nitrate	NaNO ₃	
Potassium nitrate	KNO ₃	

- 6**
- 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:
- the largest percentage of nitrogen?
 - 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:
- $$\text{NH}_3(\text{aq}) + \text{HNO}_3(\text{aq}) \rightarrow \text{NH}_4\text{NO}_3(\text{aq})$$
- 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:
- higher or lower temperatures
 - higher or lower pressures.
- d** What is the percentage yield of ammonia if the conditions used to run the process are:
- a temperature of 350°C and a pressure of 100 atmospheres?
 - 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.
- Industry normally requires water which has been softened.
 - Hard water is good for the promotion of healthy bones and teeth.
 - Hard water causes kettles to fur. This 'fur' can be removed by using a dilute acid.
 - Hard water wastes soap.
 - Hard water can coat lead pipes and reduce the possibility of lead poisoning.

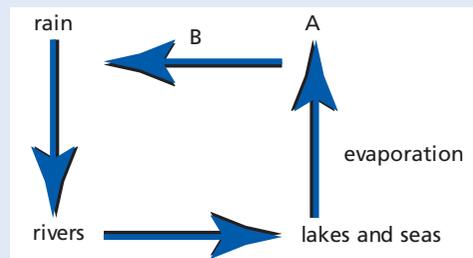
- 9 The results of testing five samples of water from different areas are shown in the table below. The soap solution was gradually added to 25 cm³ of each sample of water with shaking until a permanent lather (one which lasts for 20 seconds) was obtained.

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.

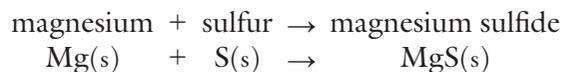
Uses of sulfur

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.

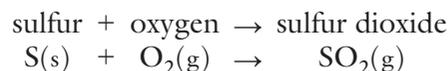
Chemical properties of sulfur

Sulfur will react with both metals and non-metals.

- It reacts with magnesium metal to form magnesium sulfide.

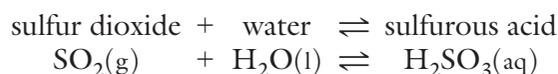


- It reacts with oxygen to produce sulfur dioxide gas.



● 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.



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

- '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.

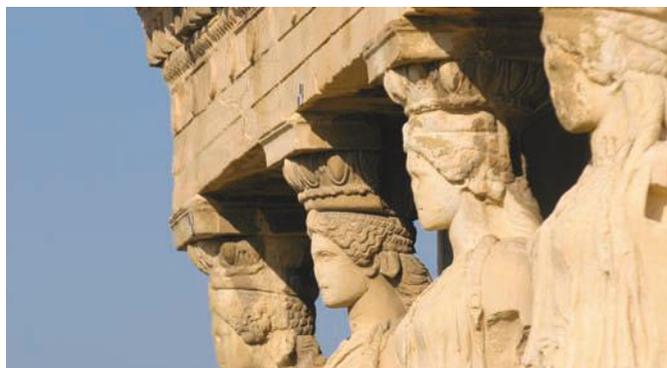
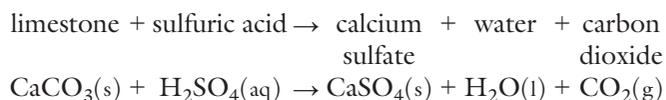


Figure 12.4 Acid rain is responsible for much of the damage to this temple on the Acropolis in Athens.

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.

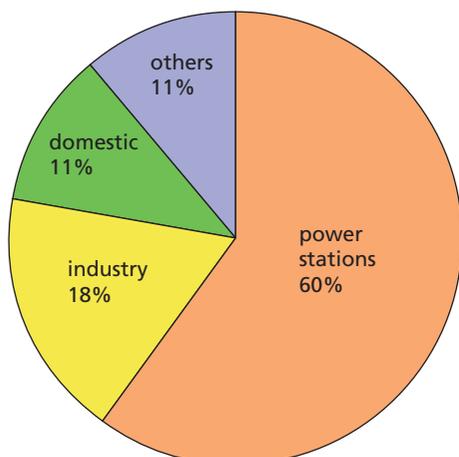


Figure 12.5 Sources of sulfur dioxide.

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).

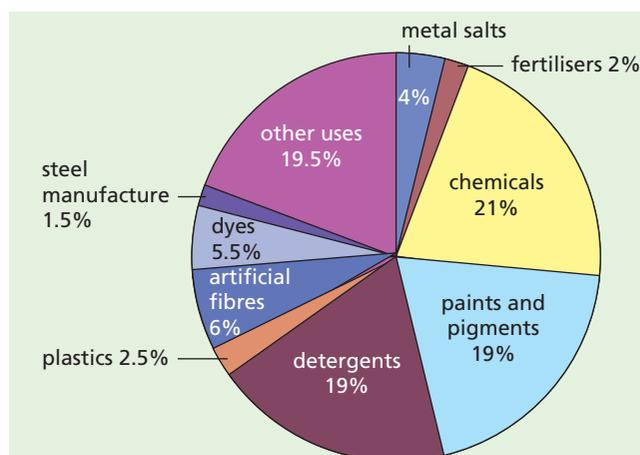


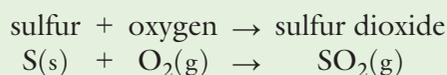
Figure 12.6 Products made from sulfuric acid include detergents, paints and pigments.



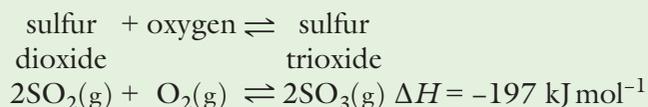
Figure 12.7 A Contact process plant used for making sulfuric acid.

The process has the following stages.

- Sulfur dioxide is first produced, primarily by the reaction of sulfur with air.



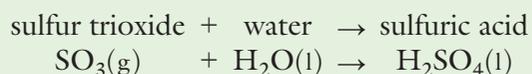
- Any dust and impurities are removed from the sulfur dioxide produced, as well as any unreacted oxygen. These 'clean' gases are heated to a temperature of approximately 450 °C and fed into a reaction vessel, where they are passed over a catalyst of vanadium(V) oxide (V₂O₅). This catalyses the reaction between sulfur dioxide and oxygen to produce sulfur trioxide (sulfur(VI) oxide, SO₃).



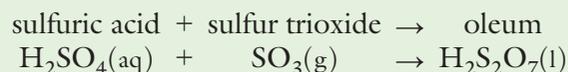
This reaction is reversible and so the ideas of Le Chatelier (Chapter 11, p. 178) can be used to increase the proportion of sulfur trioxide in the equilibrium mixture. The forward reaction is exothermic and so would be favoured by low temperatures. The temperature of 450 °C used is an optimum temperature 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.

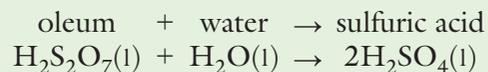
- If this sulfur trioxide is added directly to water, sulfuric acid is produced. This reaction, however, is very violent and a thick mist is produced.



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**.



The oleum formed is then added to the correct amount of water to produce sulfuric acid of the required concentration.



Questions

- 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.
- Both the following reactions are reversible:
 - $\text{X}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{XO(g)}$
 - $2\text{XO(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{XO}_2\text{(g)}$
 Suggest a reason why an increase in pressure:
 - does not favour reaction (i)
 - increases the amount of XO₂ 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.

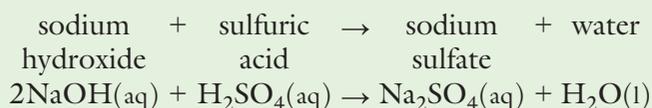
Properties of sulfuric acid

Dilute sulfuric acid

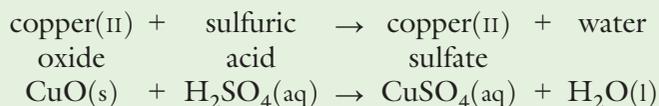
Dilute sulfuric acid is a typical strong **dibasic** acid. A dibasic acid is one with two replaceable hydrogen atoms which may produce two series of salts – normal and acid salts (Chapter 8, pp. 122 and 126).

It will react with bases such as sodium hydroxide and copper(II) oxide to produce normal salts, called **sulfates**, and water.

- With sodium hydroxide:



- With copper(II) oxide:



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).

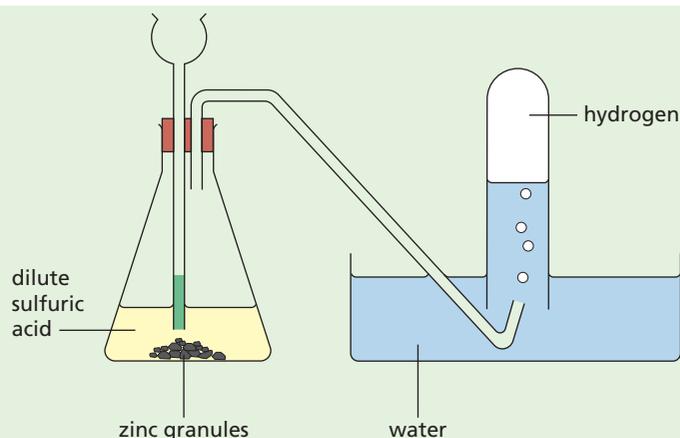
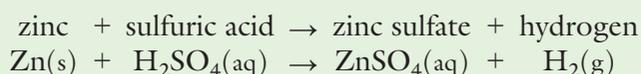
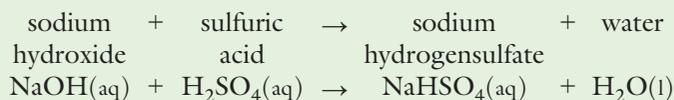


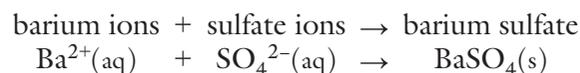
Figure 12.8 The laboratory preparation of hydrogen gas.

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 cm³ of dilute sulfuric acid were required to form the normal salt from a given volume of alkali of a particular concentration then 50 cm³ of the same acid solution would be required to produce the acid salt, sodium hydrogensulfate, from the same volume of alkali.



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.

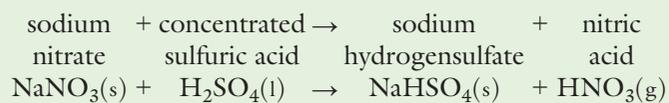


Many sulfates have very important uses, as can be seen from Table 12.1.

Table 12.1 Uses of some metal sulfates.

Salt	Formula	Use
Ammonium sulfate	(NH ₄) ₂ SO ₄	Fertiliser
Barium sulfate	BaSO ₄	'Barium meal' used in diagnostic medical X-ray studies
Calcium sulfate	CaSO ₄ · ½H ₂ O	'Plaster of Paris' used to set bones
Magnesium sulfate	MgSO ₄	In medicine it is used as a laxative

If concentrated sulfuric acid is heated with a metal nitrate such as sodium nitrate, nitric acid vapour is produced.



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 ($\text{SO}_4^{2-}(\text{aq})$), 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 ammonia 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:



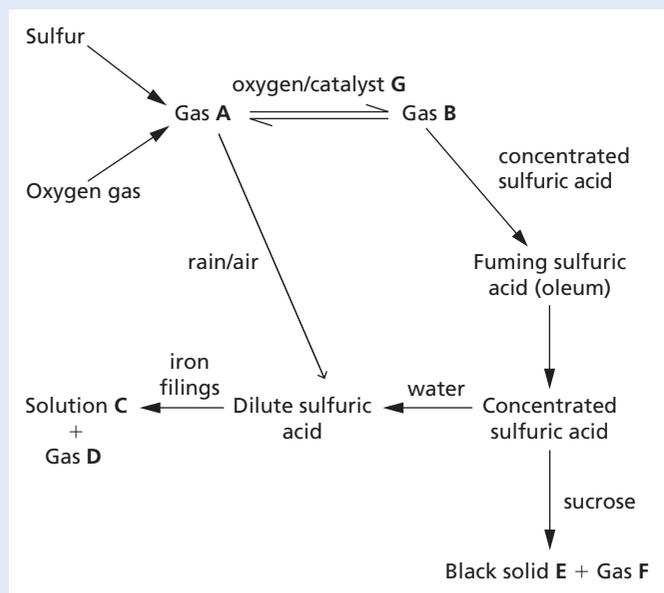
- How many electrons, protons and neutrons does a sulfur atom contain?
 - How many electrons will there be in the outer shell (energy level) of this atom?
 - To which group of the Periodic Table does sulfur belong?
 - Write down the names of two other elements found in the same group as sulfur.
 - Write down the formula of the ion that sulfur will form in reactions with metals such as magnesium.
 - Sulfur will react with powdered calcium when the mixture is heated. Write a word and a balanced chemical equation for this reaction.
 - Give two large-scale uses of sulfur.
- 2 Explain the following.
- Chemical plants that produce sulfuric acid are often located on the coast.
 - Even though more sulfuric acid could be produced using high pressures, normal atmospheric pressure is used.
 - Natural rubber cannot be used to produce car tyres but vulcanised rubber can.
 - Sulfur dioxide gas is regarded as a pollutant.
 - 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.
- Write an equation for the formation of sulfur dioxide gas when this coal is burned.
 - What mass of sulfur is contained in 1500 tonnes of coal?
 - What mass of sulfur dioxide gas would be formed if 1500 tonnes of coal were burned?
 - 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

- Using the figures in the table, produce a bar chart to show the amount of sulfur dioxide produced by each of the countries listed.
- What percentage of the world's sulfur dioxide is produced in:
 - France?
 - North America?
- 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.
- If the total amount of sulfur dioxide produced by Canada is 5 million tonnes per year, what amount is produced by:
 - power stations?
 - domestic users?
 - industry?

5 Study the following reaction scheme:



- Identify the substances **A** to **F** by giving their names and formulae.
- Write a balanced chemical equation for the formation of gas **B**.
- (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**.
- How would you obtain solid **C** from the solution **C**?
- Which pathway shows the formation of acid rain?
- In which way is the concentrated sulfuric acid acting in its reaction with sucrose?
- Where does the oxygen gas come from to form gas **A**?

- 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.
 - Write a balanced chemical equation for the reaction between dilute sulfuric acid and potassium hydroxide.
 - Calculate the number of moles of potassium hydroxide solution used in the reaction.
 - How many moles of dilute sulfuric acid would this number of moles of potassium hydroxide react with?
 - Calculate the concentration of the dilute sulfuric acid.
 - Which indicator could have been used to determine when neutralisation had just occurred?
- 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.
- 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.
 - How is the sulfur dioxide made? Give an equation for the reaction.
 - Write the balanced chemical equation for the reaction that takes place between sulfur dioxide and air.
 - Name a suitable catalyst for the reaction.
 - Why is a catalyst necessary for this process?
 - 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.
 - 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.

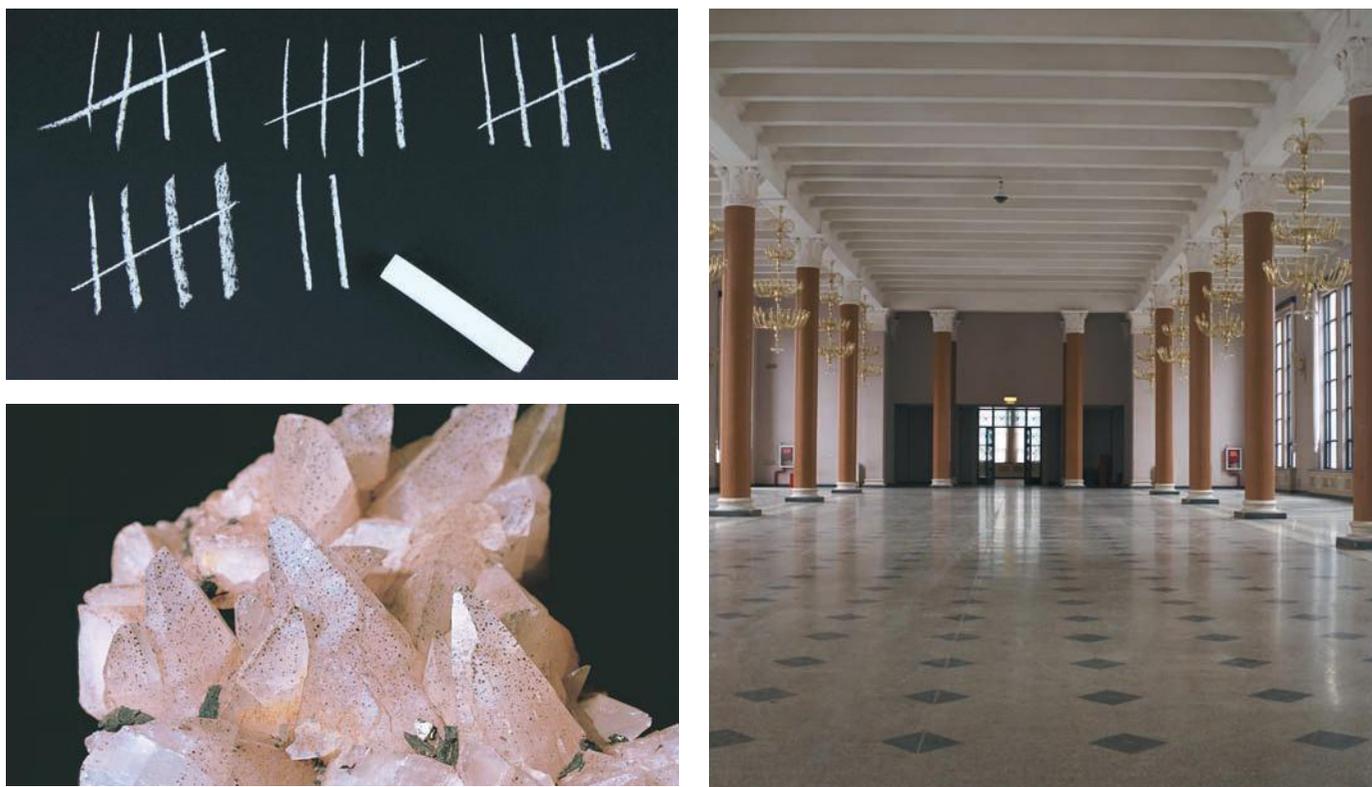


Figure 13.2 Chalk, calcite and marble are all forms of calcium carbonate.

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).

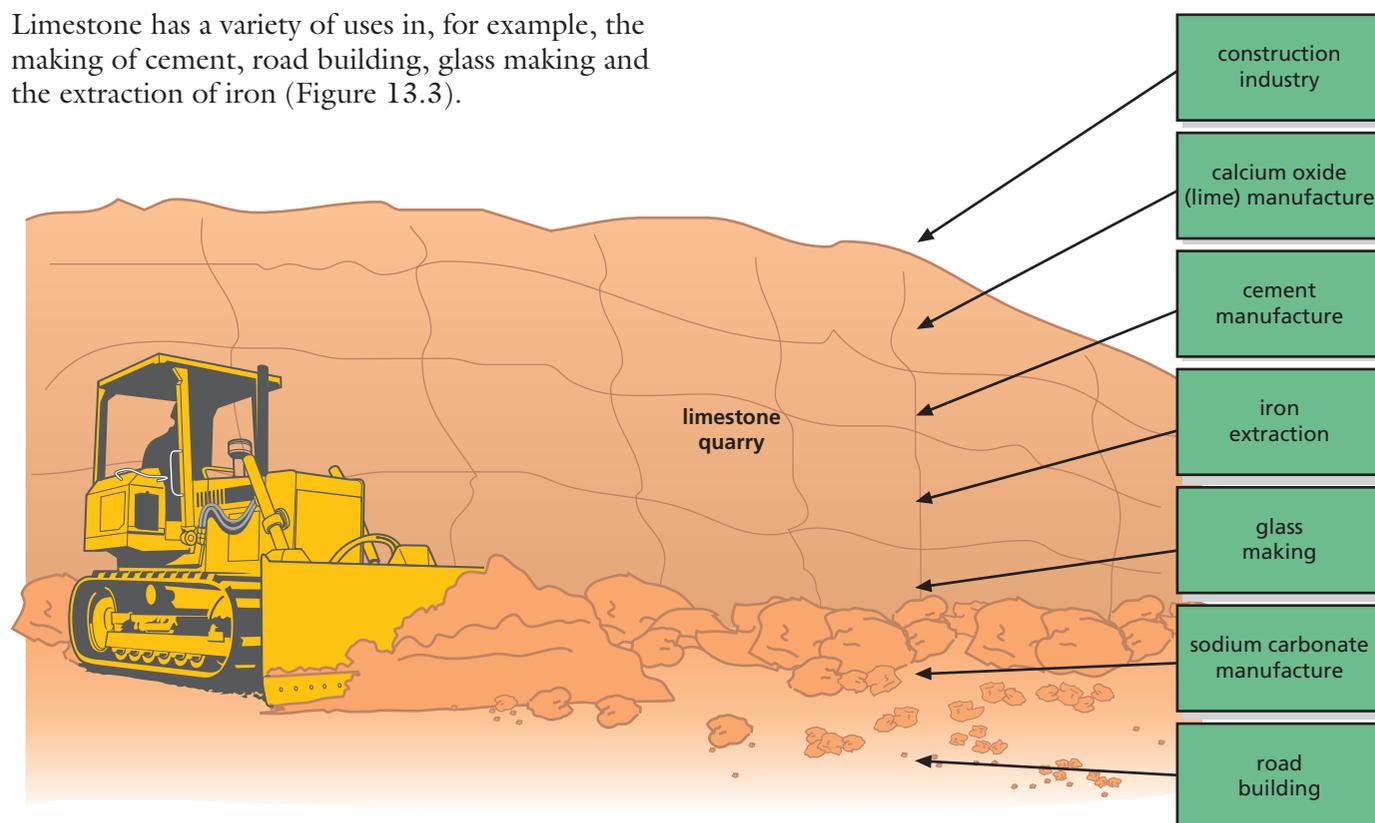


Figure 13.3 Uses of limestone.

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.

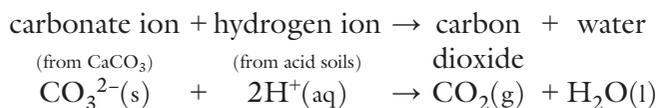


Figure 13.4 Spreading limestone on to soil.

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 ($\text{Ca}(\text{AlO}_2)_2$) 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.

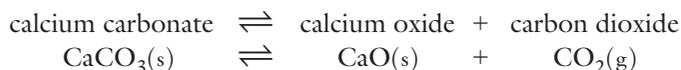


Figure 13.6 This bridge in Brazil uses reinforced concrete.

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.



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.

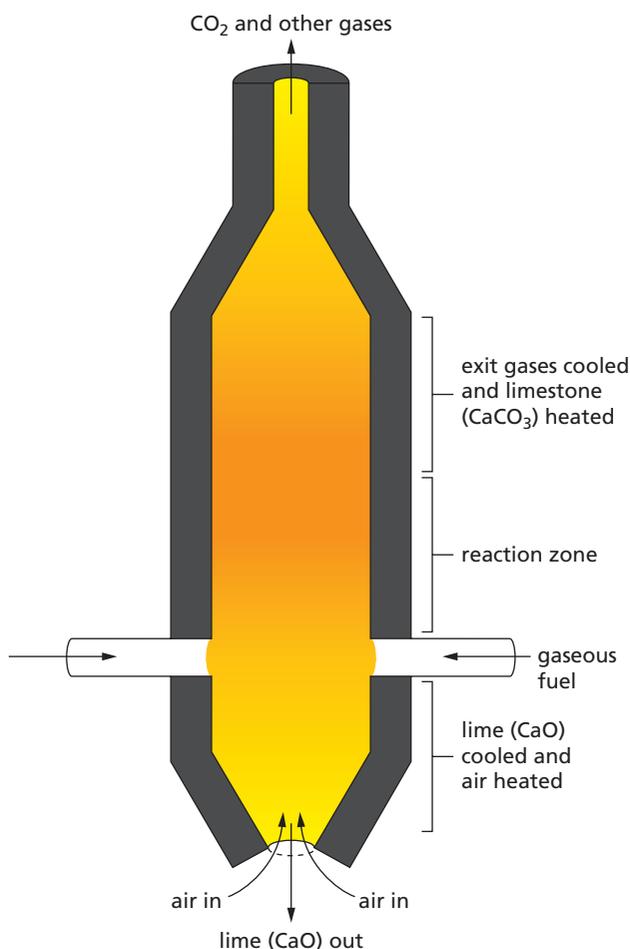


Figure 13.7 The calcium oxide produced from this process is known as quicklime or lime and is used in large quantities in the manufacture of soda glass.

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

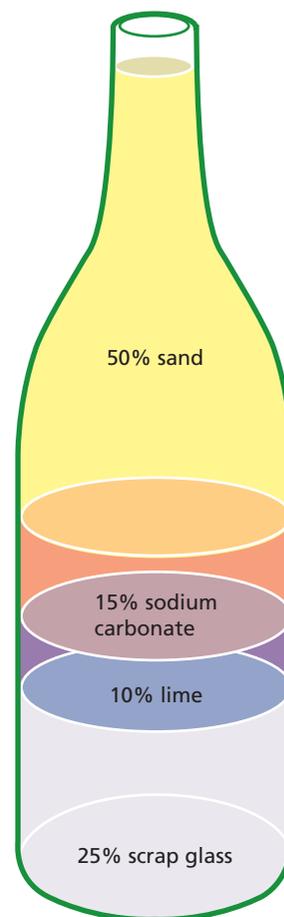


Figure 13.8 The composition of glass.

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.

Large amounts of calcium oxide are also converted into calcium hydroxide ($\text{Ca}(\text{OH})_2$) which is called **slaked lime**.

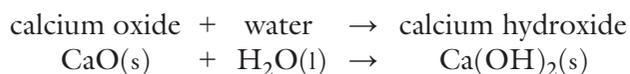
Manufacture of calcium hydroxide – slaked lime

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.

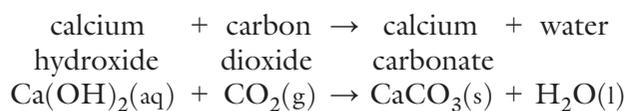


Figure 13.9 Large-scale production of calcium hydroxide.

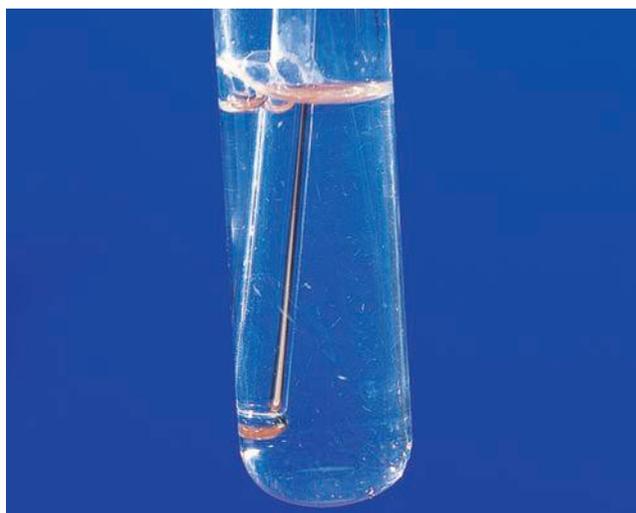
This process can be shown on the small scale in the laboratory by heating a lump of limestone very strongly to convert it to calcium oxide. Water can then be carefully added dropwise to the calcium oxide. An exothermic reaction takes place as the water and calcium oxide react together in this slaking process to form calcium hydroxide.



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):



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.

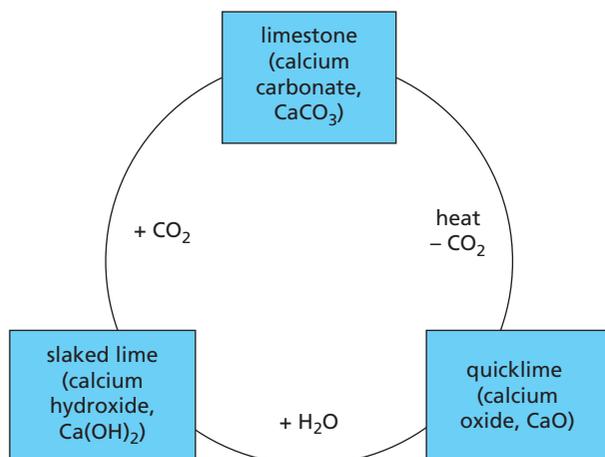
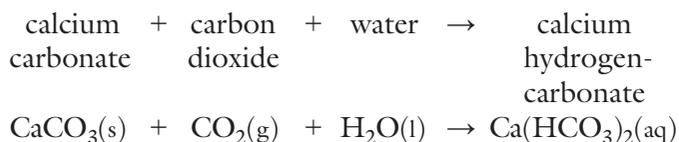
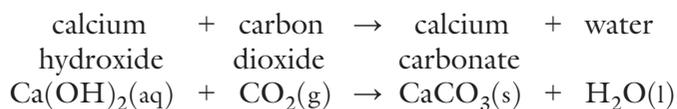


Figure 13.11 The limestone cycle.

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 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.



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 (H₂CO₃) and contain the carbonate ion (CO₃²⁻). Many of them occur naturally in rock formations. For example, in

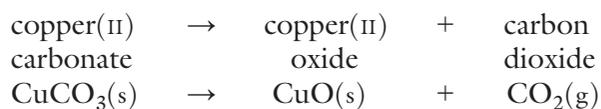


Figure 13.12 Three naturally occurring carbonates – dolomite with magnesite in front (left) and malachite (right).

addition to the calcium carbonate in limestone, chalk and marble, malachite is copper(II) carbonate and dolomite is magnesium carbonate (Figure 13.12).

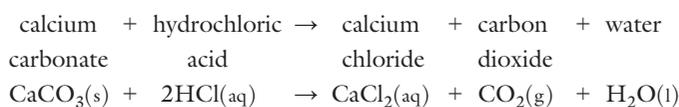
Properties of carbonates

- Most metal carbonates thermally decompose when heated to form the metal oxide and carbon dioxide gas (see Chapter 10, p. 152). For example,



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.



- 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

- Write word and balanced chemical equations to show the effect of heat on cobalt carbonate (CoCO_3) and nickel carbonate (NiCO_3).
- Write word and balanced chemical equations for the reaction of dilute hydrochloric acid with:
 - copper(II) carbonate
 - 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%.

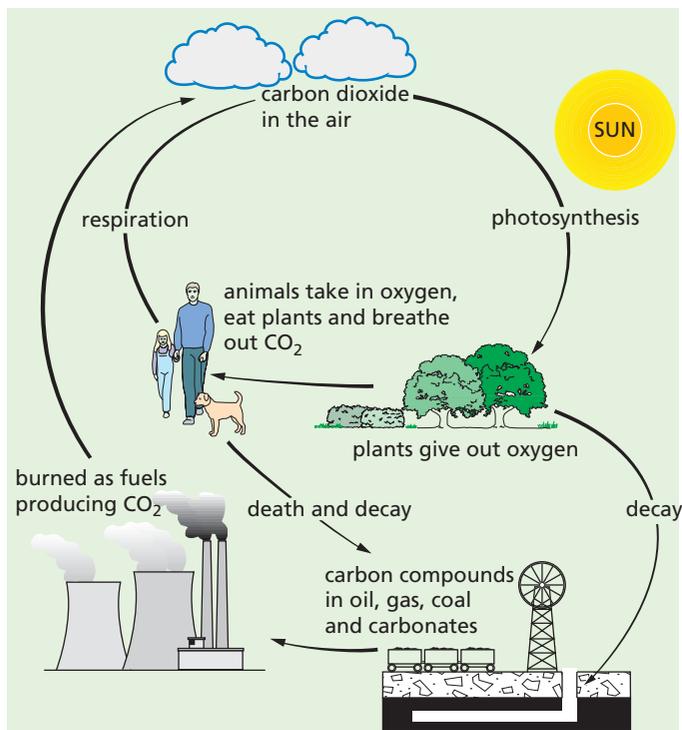
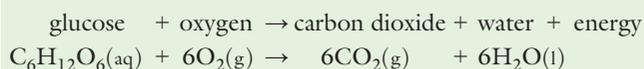
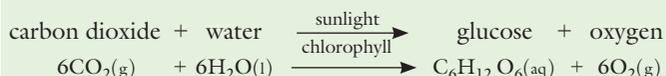


Figure 13.13 The carbon cycle.

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.



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.



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.

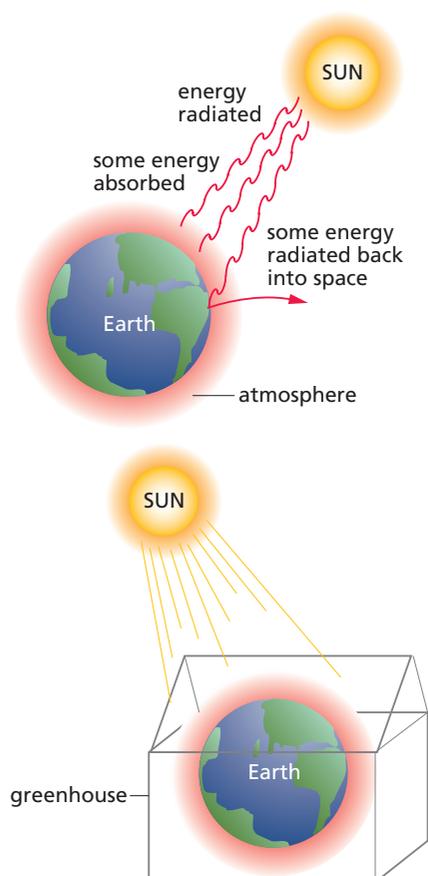


Figure 13.14 The greenhouse effect.

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.

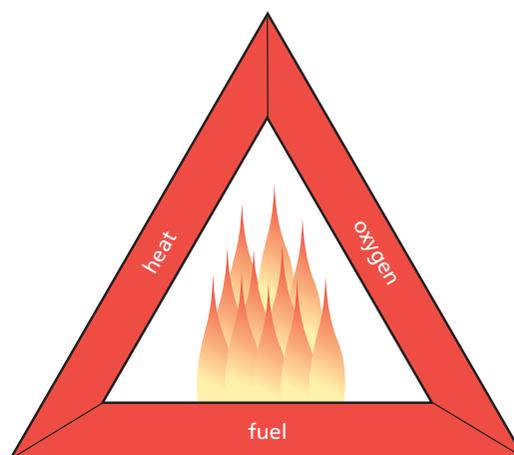
Questions

- '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.
- Use your research skills to find out about:
 - the carbon trading scheme
 - low-carbon homes
 - the Kyoto Protocol.
- 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.



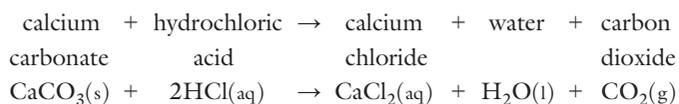
b Carbon dioxide fire extinguishers starve the fire of oxygen. They are used mainly for electrical fires.

Figure 13.15

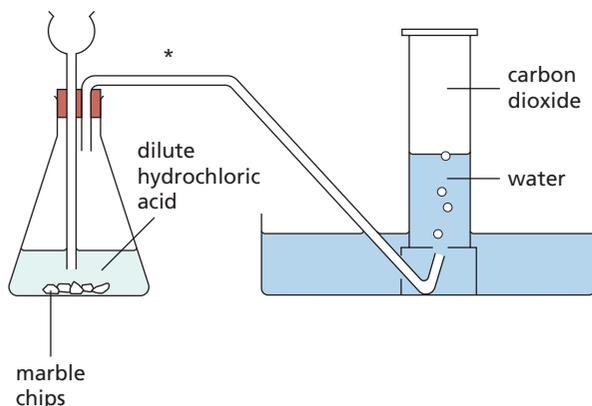
- **Refrigerants.** Solid carbon dioxide (dry ice) is used for refrigerating ice cream, meat and soft fruits. It is used for this purpose because it is colder than ice and it sublimates (p. 5), and so it does not pass through a potentially damaging liquid stage.
- **Special effects.** Carbon dioxide 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).



A suitable apparatus for preparing carbon dioxide is shown in Figure 13.16.



* If dry gas is required, then it is passed through concentrated sulfuric acid (to dry it) and then collected as shown below

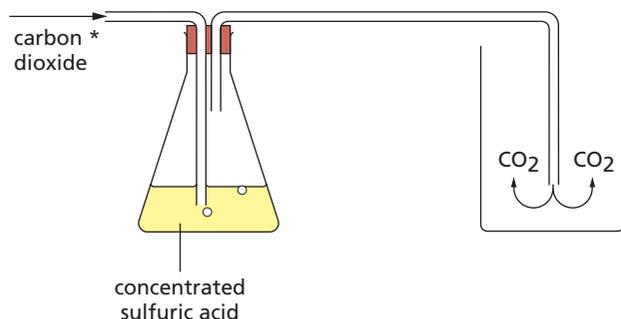


Figure 13.16 Preparation and collection of carbon dioxide gas.

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

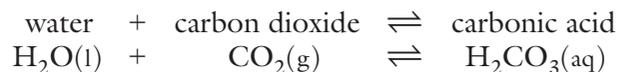
Physical properties

Carbon dioxide is:

- a colourless gas
- sparingly soluble in water
- denser than air.

Chemical properties

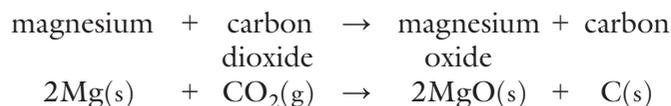
- When bubbled into water it dissolves slightly and some of the carbon dioxide reacts, forming a solution of the weak acid carbonic acid which shows a pH of 4 or 5.



- 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).

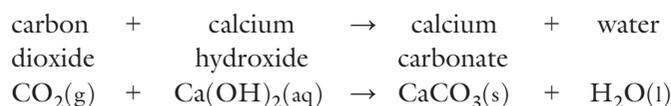


Figure 13.17 When magnesium burns in carbon dioxide gas, magnesium oxide (white) and carbon (black) are produced.

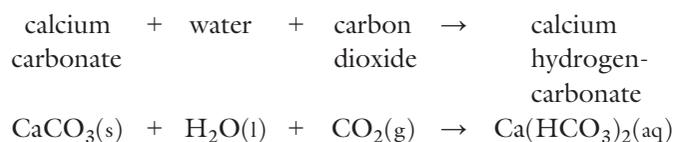


This reaction is a good example of a **redox** process (Chapter 2, p. 14). Which of these substances has been oxidised and which has been reduced? Which of these substances is the oxidising agent and which is the reducing agent?

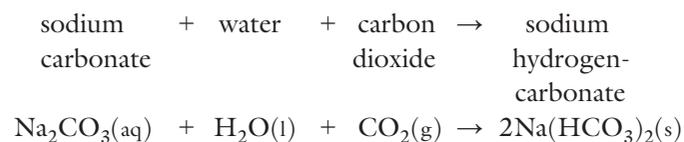
- When carbon dioxide is bubbled through limewater (calcium hydroxide solution), a white precipitate is formed. This white solid is calcium carbonate (CaCO_3). This reaction is used as a test to show that a gas is carbon dioxide.



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.



- 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 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

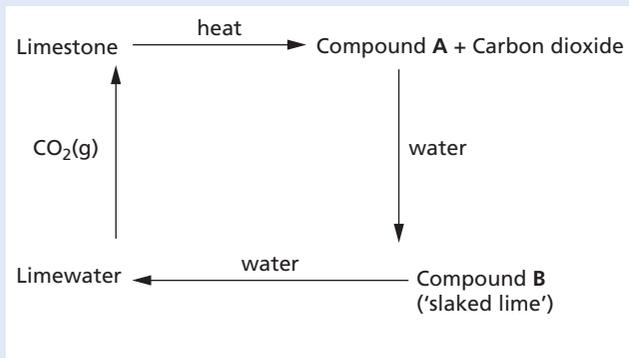
After studying Chapter 13 you should know and understand the following terms.

- **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.

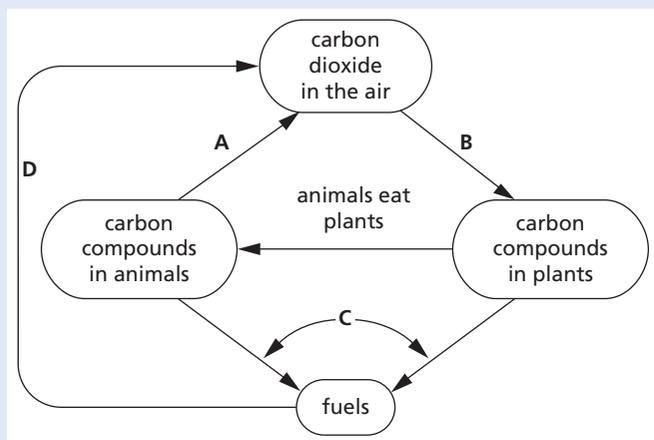
Inorganic carbon chemistry

Additional questions

1 This question is about the limestone cycle.



- Name and give the formula of:
 - compound A
 - compound B.
 - Write balanced chemical equations for the formation of both compounds A and B.
 - Name and give the symbol/formula for the ions present in limewater.
 - 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.

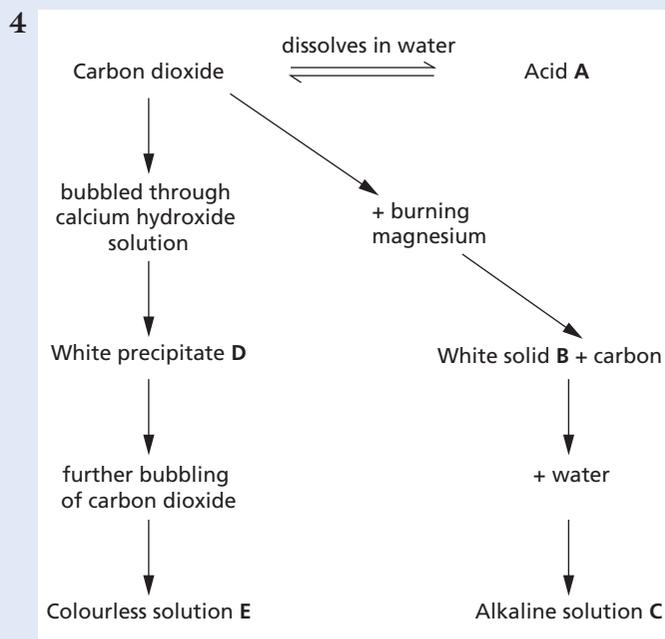


- Name the processes A, B, C and D.
- Write balanced chemical equations to represent the processes taking place in A and B.
- There are a number of fuels which could be placed in the 'fuels' box. Name three such fuels.
- 'The air contains approximately 0.03% of carbon dioxide and this amount is almost constant.' Explain why this is a true statement.

- What is the 'greenhouse effect'?
- Discuss the possible consequences of the greenhouse effect if nothing is done to counteract it.

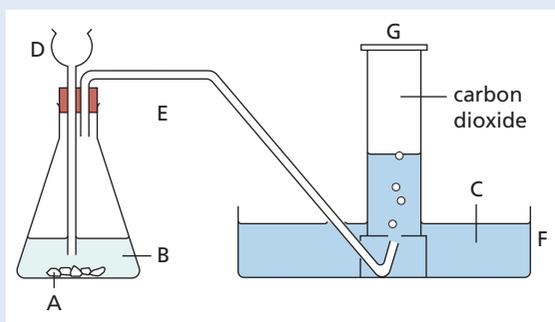
3 Limestone is an important raw material used in many different industries.

- One of the properties of limestone is that it reacts with acids.
 - Why do farmers spread powdered limestone on their fields?
 - How can buildings made of limestone be affected by 'acid rain'?
 - Write an ionic equation which would represent the reactions taking place in both (i) and (ii).
- Limestone is used in the manufacture of iron (Chapter 10, p. 158).
 - Why is it added to the blast furnace along with coke and haematite?
 - Write chemical equations for the reactions it is involved in, in this process.
- Name a building material made by heating a mixture of limestone and clay in a rotary kiln.
 - What substance is the dry product produced in the rotary kiln added to? Explain why this substance is added.



- 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 dm³ 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** **CaCO₃** **acids**
thermal **CO₂** **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:
- $$\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$$
- 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.

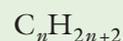
You saw in Chapter 6 that crude oil is made up of a complex mixture of hydrocarbon compounds. This mixture is a very important raw material in the organic chemical industry. Our world would be a very different place without these substances. Can you imagine life without the substances from oil and natural 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:

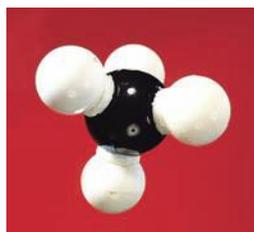
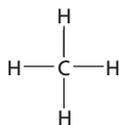


where n is the number of carbon atoms present.

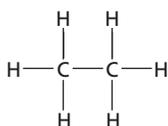
As you go up a homologous series, in order of increasing number of carbon atoms, the physical properties of the compounds gradually change. For example, the melting and boiling points of the alkanes shown in Table 14.1 gradually increase. This is due to an increase in the intermolecular forces (van der 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.

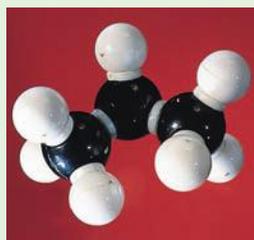
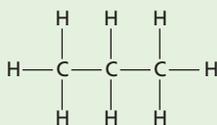
methane



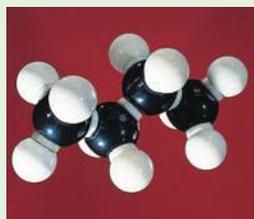
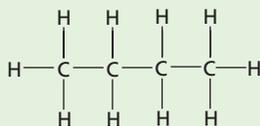
ethane



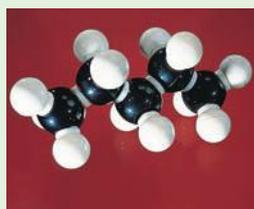
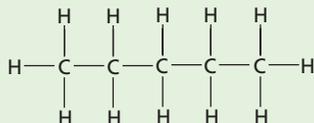
propane



butane



pentane



hexane

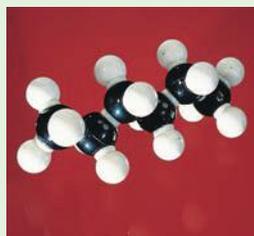
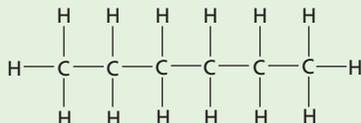


Figure 14.2 The fully displayed (or structural) formulae and molecular models of the first six alkanes.

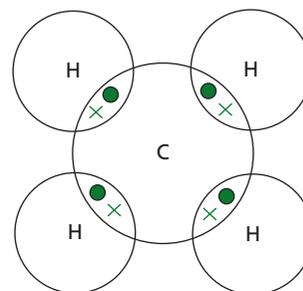
methane molecule (CH₄)

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 ₂ H ₆	-183	-89	Gas
Propane	C ₃ H ₈	-188	-42	Gas
Butane	C ₄ H ₁₀	-138	0	Gas
Pentane	C ₅ H ₁₂	-130	36	Liquid
Hexane	C ₆ H ₁₄	-95	69	Liquid

Questions

- Estimate the boiling points for the alkanes with formulae:
 - C₇H₁₆
 - C₈H₁₈
- Name the alkanes which have the following formulae:
 - C₇H₁₆
 - C₁₀H₂₂

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.

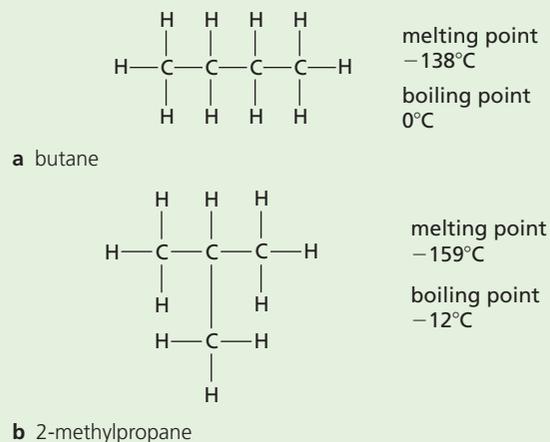


Figure 14.4 The isomers of C_4H_{10} .

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!

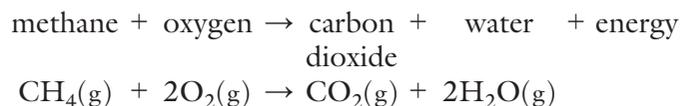
Question

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.



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.



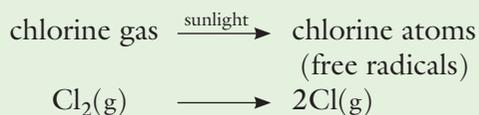
a This is burning methane.



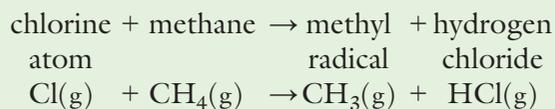
b Central heating systems can be run on propane.

Figure 14.5

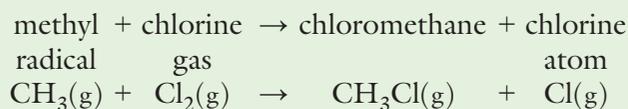
Another useful reaction worth noting is that between the alkanes and the halogens. For example, methane and chlorine react in the presence of sunlight (or ultraviolet light). The ultraviolet light splits the chlorine molecules into atoms. When this type of reaction takes place, these atoms are called **free radicals** and they are very reactive.



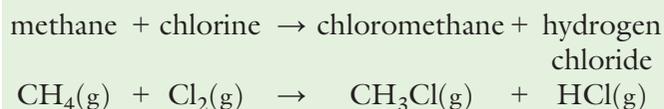
The chlorine atoms then react further with methane molecules, and a hydrogen chloride molecule is produced along with a methyl free radical.



The methyl free radical reacts further.

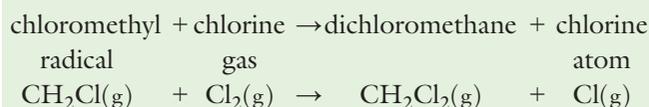
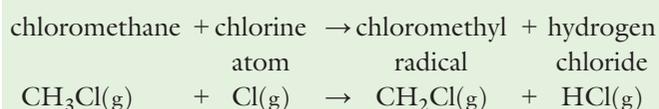


This chlorine free radical, in turn, reacts further and the process continues until all the chlorine and the methane have been used up. This type of process is known as a **chain reaction** and it is very fast. The overall chemical equation for this process is:



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.



Many of these so-called **halogenoalkanes** are used as solvents. For example, dichloromethane is used as a solvent in paint stripper (Figure 14.6).

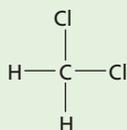


Figure 14.6 Dichloromethane is used as a solvent in paint stripper.

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

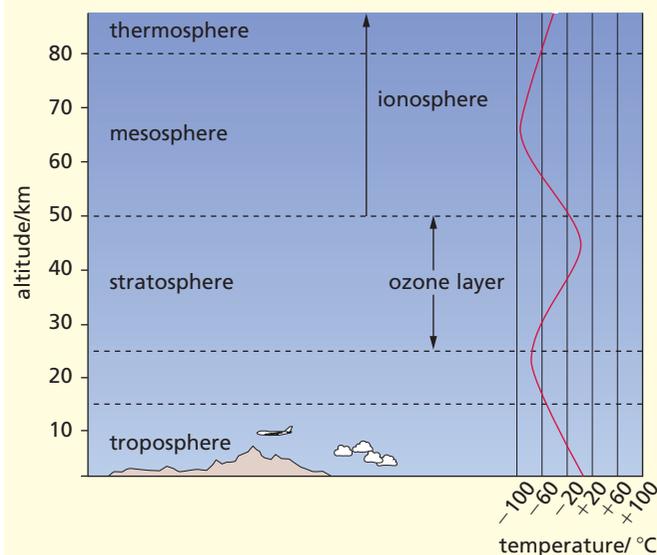
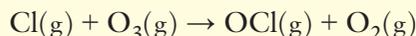


Figure 14.7 The ozone layer is between 25 km and 50 km above sea level.

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.



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

- Use your research skills to find out:
 - any other sources of methane found in nature
 - 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).

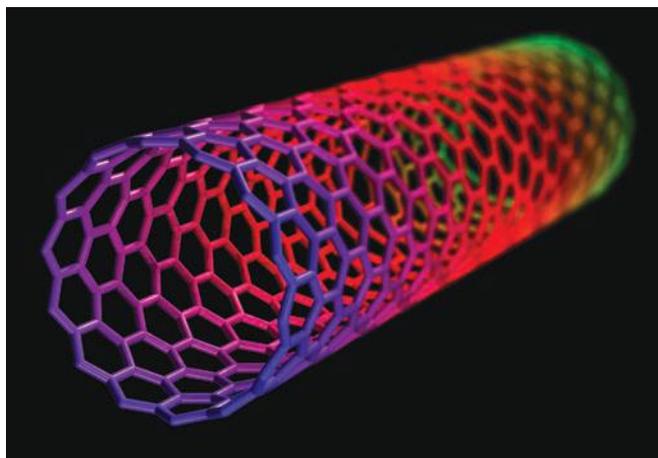


Figure 14.8 These nanotubes are being used in electrical circuitry.

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

- Write a balanced chemical equation to represent the combustion of propane.
- In what mole proportions should chlorine and methane be mixed to produce:
 - mainly chloromethane?
 - mainly tetrachloromethane?
- Describe a method you would use to separate chloromethane from the other possible reaction products when methane reacts with chlorine.
- 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

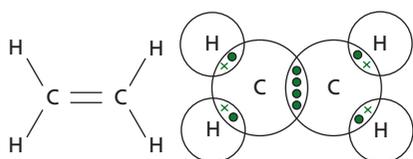


Figure 14.9 The bonding in ethene, the simplest alkene.

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.

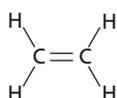
Naming the alkenes

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.

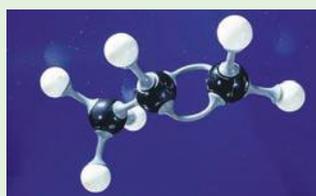
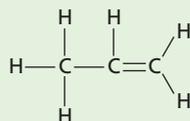
Table 14.2 The first three alkenes and their physical properties.

Alkene	Formula	Melting point/°C	Boiling point/°C	Physical state at room temperature
Ethene	C ₂ H ₄	-169	-104	Gas
Propene	C ₃ H ₆	-185	-47	Gas
Butene	C ₄ H ₈	-184	-6	Gas

ethene



propene



butene

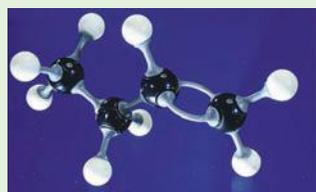
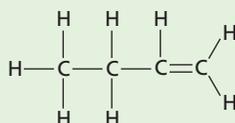
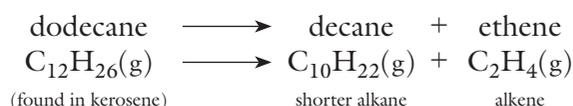


Figure 14.10 Structure and shape of the first three alkenes.

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.



Another possibility is:



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.

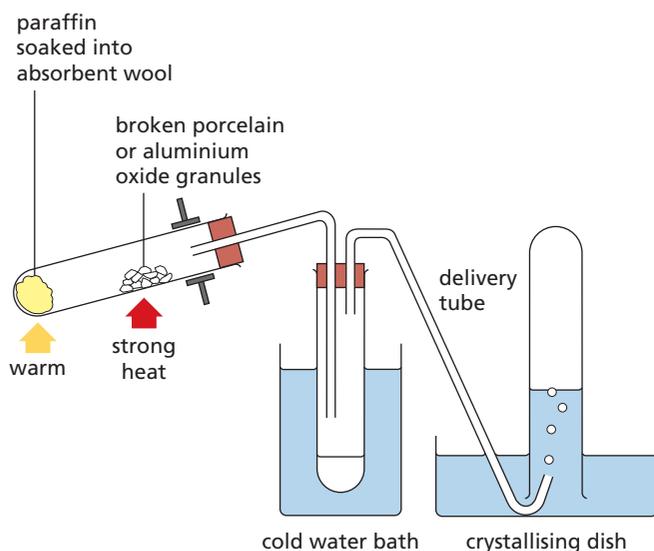


Figure 14.11 The cracking of an alkane in the laboratory.

Questions

- Using the information in Table 14.2 (p. 223), make an estimate of the boiling point of pentene.
- 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).

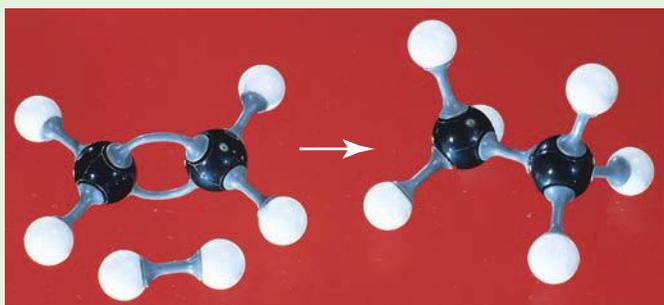
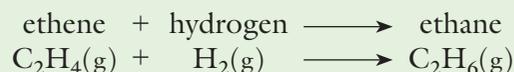


Figure 14.12 The addition of hydrogen to ethene using molecular models.

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.

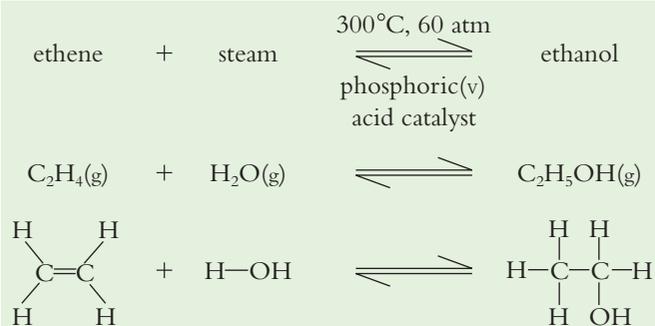


Hydrogenation reactions like the one shown with ethene are used in the manufacture of margarines from vegetable oils. Vegetable oils contain fatty acids, such as linoleic acid (C₁₈H₃₂O₂). These are unsaturated molecules, containing several double bonds. These double bonds make the molecule less flexible. 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 × 10⁵ pascals).

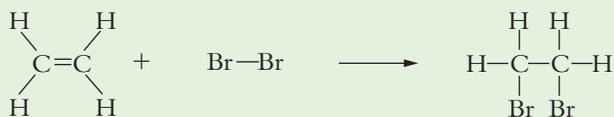
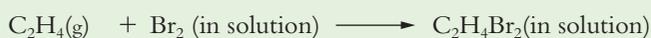
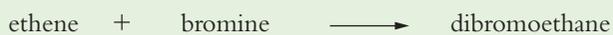


This reaction is reversible as is shown by the equilibrium (\rightleftharpoons) sign. The conditions have been chosen to ensure the highest possible yield of ethanol. In other words, the conditions have been chosen so that they favour the forward reaction.

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.



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, $\text{C}_4\text{H}_9\text{OH}$
- Octane, C_8H_{18}
- Nonane, C_9H_{20}
- Methanoic acid, HCOOH
- Butene, C_4H_8

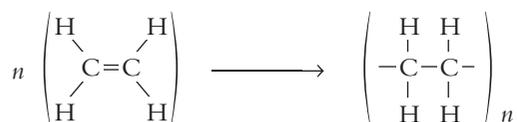
State why you have chosen your answers.



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).



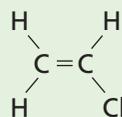
Figure 14.15 These crates are made from poly(ethene).

Other alkene molecules can also produce substances like poly(ethene); for example, propene produces poly(propene), which is used to make ropes and packaging.

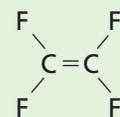
When small molecules like ethene join together to form long chains of atoms, called **polymers**, the process is called **polymerisation**. The small molecules, like ethene, which join together in this way are called **monomers**. A polymer chain, a very large molecule or a macromolecule, often consists 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.

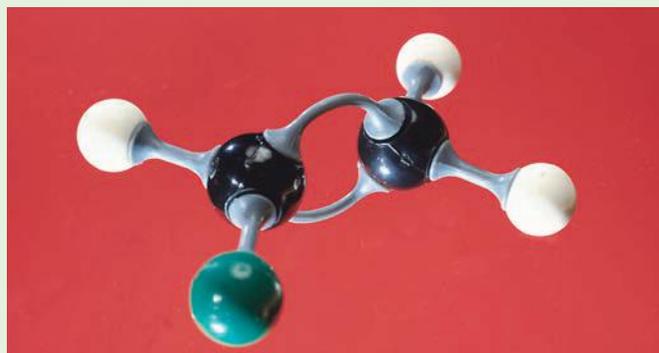
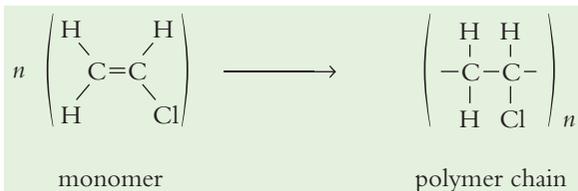


PVC monomer
(vinyl chloride or
chloroethene)

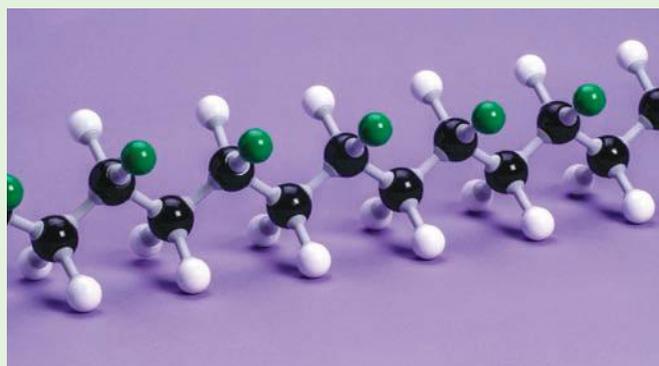


PTFE monomer
(tetrafluoroethene)

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

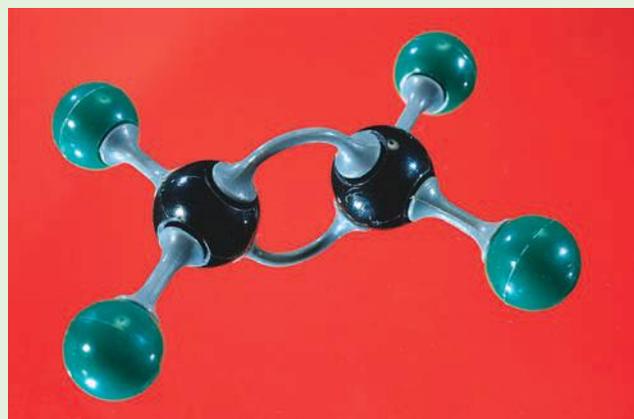


Figure 14.17 These pipes are made from PVC.

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.



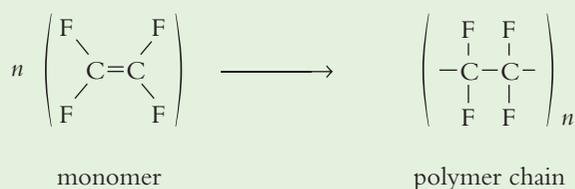
a Model of tetrafluoroethene, the PTFE monomer.



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)	$\text{CH}_2 = \text{CH}_2$	Tough, durable	Carrier bags, bowls, buckets, packaging
Poly(propene)	$\text{CH}_3\text{CH} = \text{CH}_2$	Tough, durable	Ropes, packaging
PVC	$\text{CH}_2 = \text{CHCl}$	Strong, hard (less flexible than poly(ethene))	Pipes, electrical insulation, guttering
PTFE	$\text{CF}_2 = \text{CF}_2$	Non-stick surface, withstands high temperatures	Non-stick frying pans, soles of irons
Polystyrene	$\text{CH}_2 = \text{CHC}_6\text{H}_5$	Light, poor conductor of heat	Insulation, packaging (especially as foam)
Perspex	$\text{CH}_2 = \text{C}(\text{CO}_2\text{CH}_3)\text{CH}_3$	Transparent	Used as a glass substitute

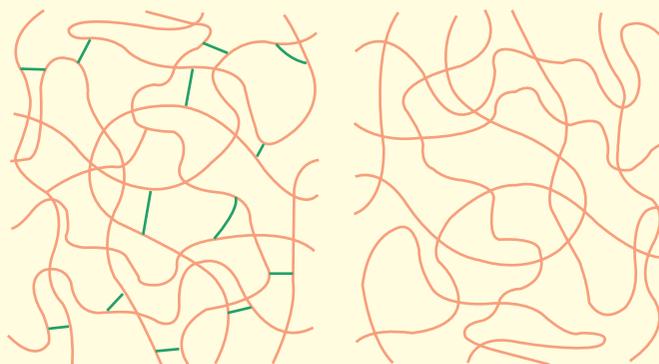
Thermosoftening and thermosetting plastics

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.

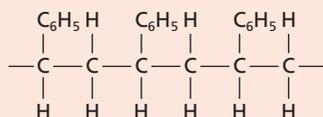


b This plastic dissolves in water.

Figure 14.24

Questions

- Write the general equation to represent the formation of polystyrene from its monomer.
- Give two advantages and two disadvantages of plastic waste (rubbish).
- Draw the structure of the repeating unit of the addition polymer formed from $\text{CH}_3\text{—CH=CH}_2$.
- Draw the structure of the monomer from which the addition polymer below has been produced.



Checklist

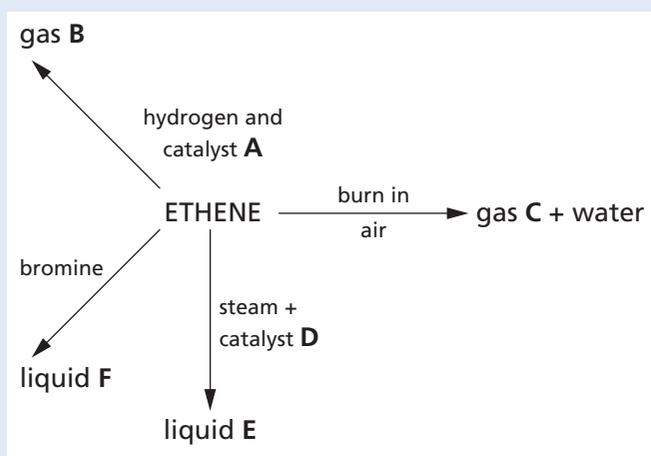
After studying Chapter 14 you should know and understand the following terms.

- **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 $\text{C}_n\text{H}_{2n+2}$. 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, $800\text{--}850^\circ\text{C}$.
- **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.

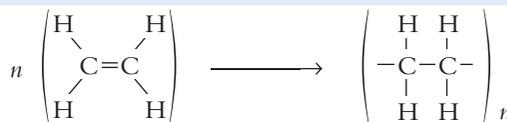
Organic chemistry 1

Additional questions

- 1 Explain the following.
- Ethene is called an unsaturated hydrocarbon.
 - The cracking of larger alkanes into simple alkanes and alkenes is important to the petrochemical industry.
 - 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.



- Give the names and formulae for substances A to F.
 - Write a word and balanced chemical equation to represent the reaction in which liquid E is formed.
 - What reaction conditions are required for the process to take place?
 - Hydrogen is used in the production of margarine to remove unsaturation. Explain what you understand by this statement.
 - Name the homologous series that gas B belongs to.
 - 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.
- 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.
 - Write a word and balanced chemical equation for the reaction between methane and chlorine.
 - Name the type of reaction taking place.
 - Highly reactive chlorine atoms are produced in the presence of ultraviolet light. When atoms are produced in this way, what are they called?
 - Write a balanced chemical equation for the reaction which takes place between CHF_3 and Cl_2 to produce a chlorofluorocarbon (CFC).
 - Why are CFCs such a problem?
- 4 a Ethene, C_2H_4 , is the starting material for making plastic carrier bags.



- Name the type of chemical change taking place in the diagram above.
 - Name the product formed by this reaction.
 - 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.
- Explain why carrier bags should not just be thrown away.
 - Explain why the majority of plastic carrier bags are recycled.
 - 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.
- Methane, CH_4 .
 - Propene, C_3H_6 .
 - Propane, C_3H_8 .
 - 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/ $^{\circ}\text{C}$
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.

- Which fraction would contain the most volatile substances?
 - Which of the fractions would collect at the bottom of the fractionating column?
 - Which fraction could be used as a fuel for cars?
 - Which fraction would contain the largest molecules?
- b Some of the fractions undergo a further process called cracking to produce further substances.
- Explain what you understand by the term 'cracking'. What conditions are employed when cracking occurs?
 - Write a word and balanced chemical equation to show how octane can be produced by the cracking of $\text{C}_{15}\text{H}_{32}$.
- 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} , $\text{C}_{12}\text{H}_{24}$, $\text{C}_{20}\text{H}_{42}$, 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 alkanolic 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 ₃ OH	–94	64
Ethanol	CH ₃ CH ₂ OH	–117	78
Propanol	CH ₃ CH ₂ CH ₂ OH	–126	97
Butanol	CH ₃ CH ₂ CH ₂ CH ₂ OH	–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

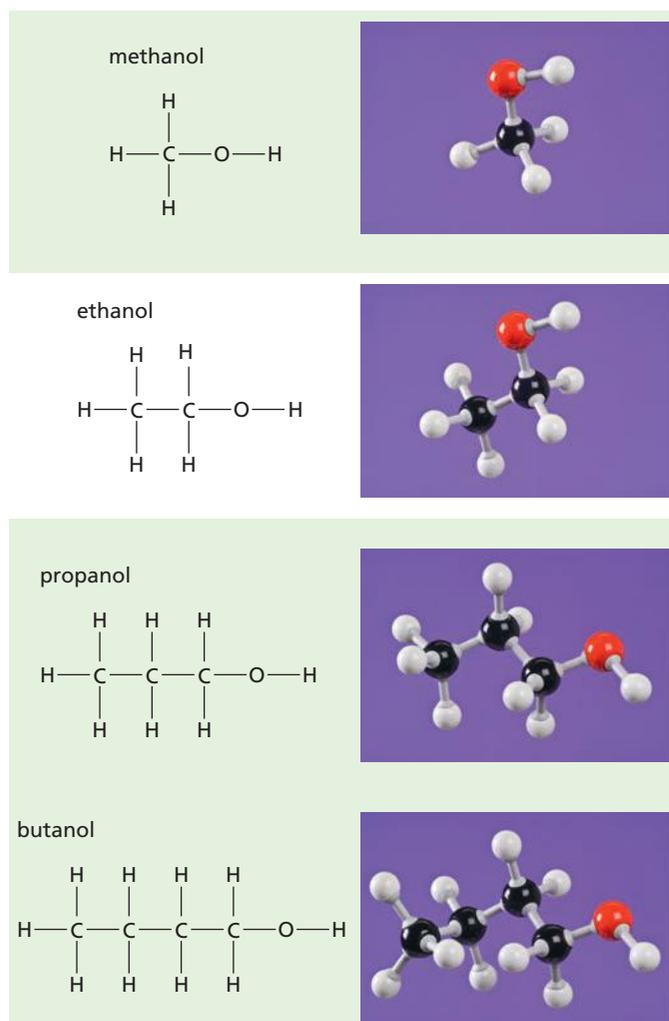


Figure 15.2 The molecules look like their corresponding models in the photographs.

number to the position of the $-\text{OH}$ group. If the $-\text{OH}$ group is positioned at the end of the alcohol, for example in $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$, 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 ($\text{H}-\text{OH}$) in that they are polar (see Chapter 11, p. 186).

Alcohol molecules are polar because of the presence of the $-\text{OH}$ 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).

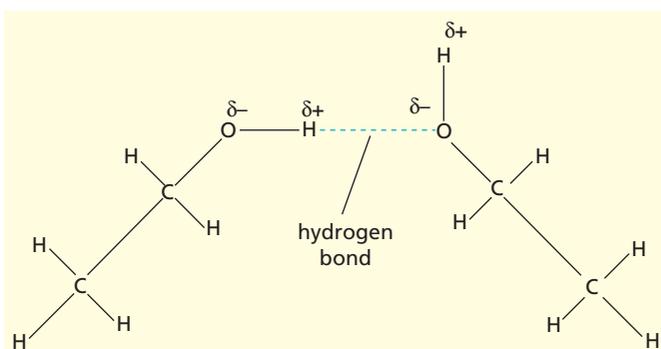


Figure 15.3 Polar $-\text{OH}$ groups in ethanol molecules lead to hydrogen bonding and an attraction between neighbouring molecules.

Because of the presence of the polar $-\text{OH}$ 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 ($\text{CH}_3\text{CH}_2\text{CH}_3$) 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 $-\text{C}_2\text{H}_5$ in $\text{C}_2\text{H}_5\text{OH}$. 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 $-\text{OH}$ 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).

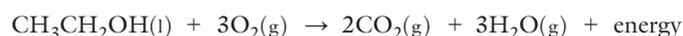
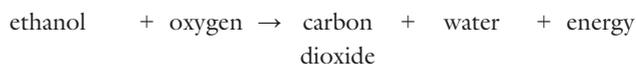


Figure 15.4 This aftershave contains alcohol.

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

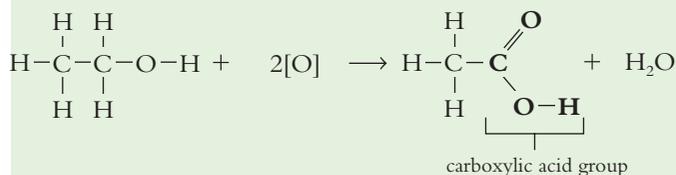
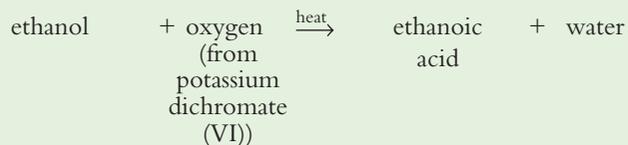
Ethanol burns quite readily with a clean, hot flame.



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.



Question

- Write a word and balanced chemical equation for:
 - the combustion of butanol
 - 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.

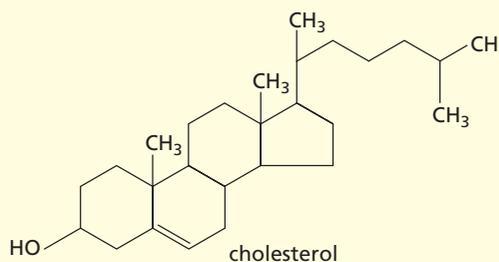
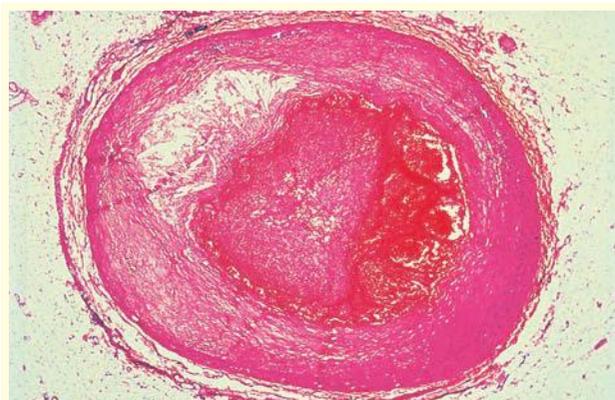
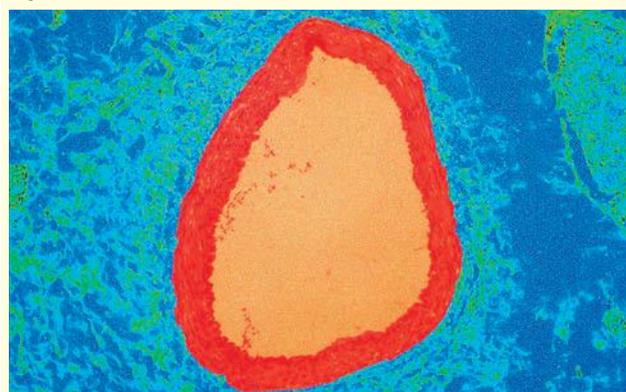


Figure 15.6 The structure of cholesterol.



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.

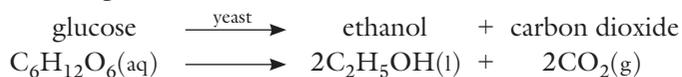


Figure 15.8 shows a simple apparatus for obtaining ethanol from glucose in the laboratory.

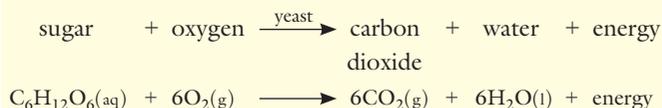
If a bottle of wine is left exposed to the atmosphere then further oxidation takes place and the ethanol is oxidised to ethanoic acid.



Figure 15.8 Fermenting glucose and yeast to produce ethanol. The bag is inflated during the experiment by CO₂.

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.



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.

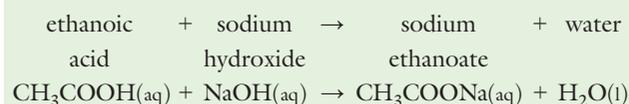
Carboxylic acids

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.

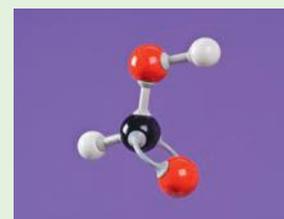
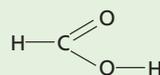
Table 15.3 Some members of the carboxylic acid series.

Carboxylic acid	Formula	Melting point/ °C	Boiling point/ °C
Methanoic acid	HCOOH	9	101
Ethanoic acid	CH ₃ COOH	17	118
Propanoic acid	CH ₃ CH ₂ COOH	-21	141
Butanoic acid	CH ₃ CH ₂ CH ₂ COOH	-6	164

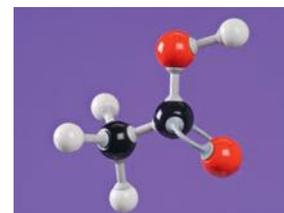
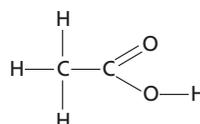
Methanoic acid is present in stinging nettles and ant stings. Ethanoic acid, however, is the most well known as it is the main constituent of vinegar. Like other acids, ethanoic acid affects indicators and will react with metals such as magnesium. However, whereas the mineral acids such as hydrochloric 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.



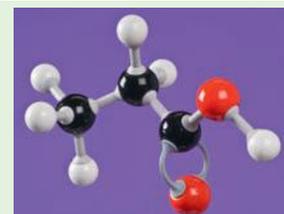
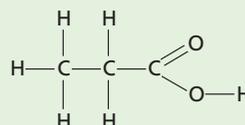
methanoic acid



ethanoic acid



propanoic acid



butanoic acid

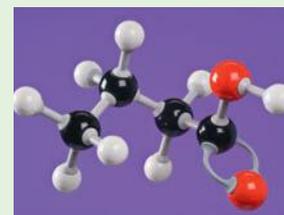
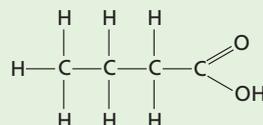
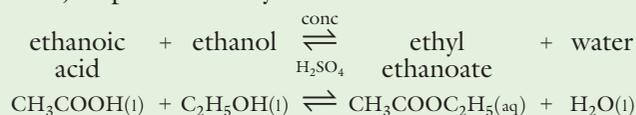


Figure 15.9 The molecules look like the 3D models in the photographs.

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.



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. $\text{CH}_3\text{COOC}_3\text{H}_7$

Ester	Made from		Structure	3D model
	Alcohol	Carboxylic acid		
Ethyl ethanoate $\text{CH}_3\text{COOC}_2\text{H}_5$	Ethanol $\text{C}_2\text{H}_5\text{OH}$	Ethanoic acid CH_3COOH		
Propyl methanoate HCOOC_3H_7	Propan-1-ol $\text{C}_3\text{H}_7\text{OH}$	Methanoic acid HCOOH		
Methyl butanoate $\text{C}_3\text{H}_7\text{COOCH}_3$	Methanol CH_3OH	Butanoic acid $\text{C}_3\text{H}_7\text{COOH}$		

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

- Write the structural formula for propanoic acid.
- 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, $\text{C}_6\text{H}_{13}\text{COOH}$
- Butanol, $\text{C}_4\text{H}_9\text{OH}$
- Octane, C_8H_{18}
- Nonane, C_9H_{20}
- Methanoic acid, HCOOH
- Pentene, C_5H_{10}
- Hexanol, $\text{C}_6\text{H}_{13}\text{OH}$

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.

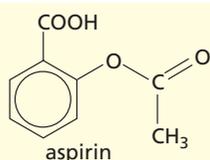
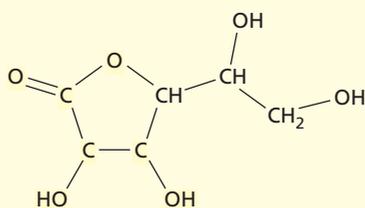


Figure 15.12 The structure of aspirin.

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.



ascorbic acid

Figure 15.13 The structure of ascorbic acid (vitamin C).

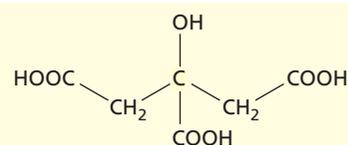
Vitamin C is important to all animals, including humans, and is vital in the production of collagen. Collagen is important in the formation of connective tissues that give our body shape and help to support vital organs. Vitamin C prevents the disease scurvy. It is found in citrus fruits and brightly 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.



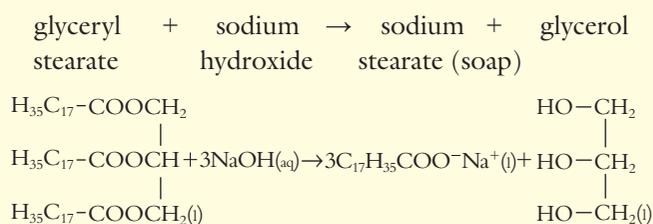
citric acid

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).



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).

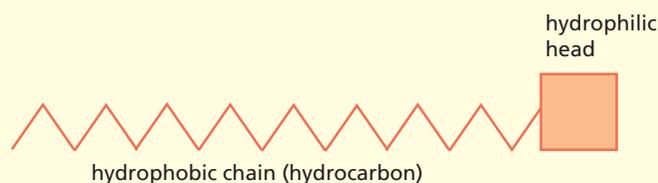


Figure 15.16 Simplified diagram of a soap molecule.

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).

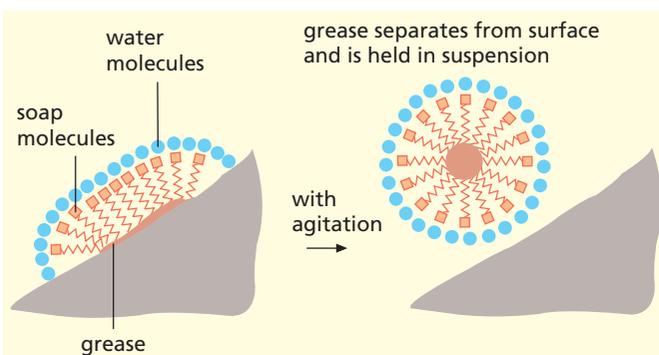


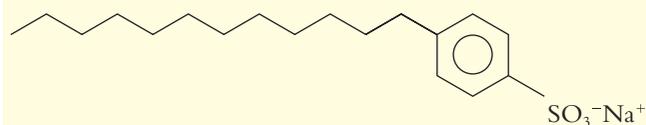
Figure 15.17 Soap dissolves grease like this.

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, $\text{C}_{18}\text{H}_{29}\text{SO}_3\text{Na}$, is given below.

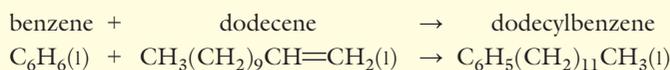


The calcium and magnesium salts of this detergent molecule are water soluble, so the problem of scum is solved. Very many of our washing powders (and liquids) contain this type of substance.

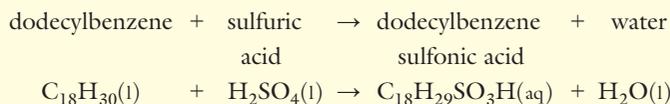
Manufacture of soapless detergents

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.

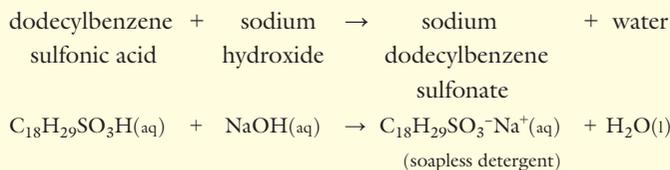
The general process involves, initially, the reaction of a long, straight-chain alkene, such as dodecene ($\text{CH}_3(\text{CH}_2)_9\text{CH}=\text{CH}_2$), with benzene.



The molecular formula of dodecylbenzene is $\text{C}_{18}\text{H}_{30}$. This compound is then reacted with concentrated sulfuric acid to give a compound that is known as a sulfonic acid.



Finally, this is reacted with the alkali sodium hydroxide, NaOH.



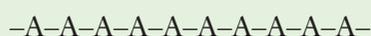
These days, soapless detergents such as this are to be found in most washing powders and liquids.

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?

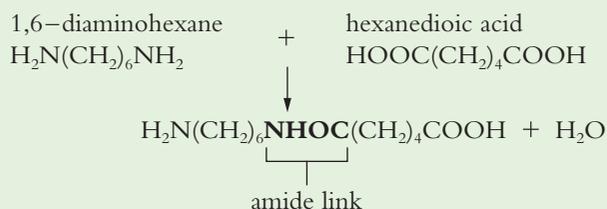
● Condensation polymers

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:



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.



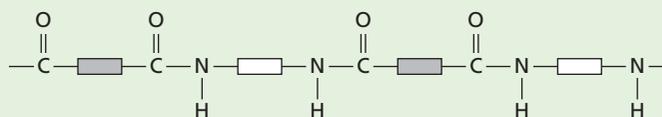
Figure 15.18 A nylon polymer chain is made up from the two molecules arranged alternately just like the two different coloured poppet beads in the photo.

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:



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.

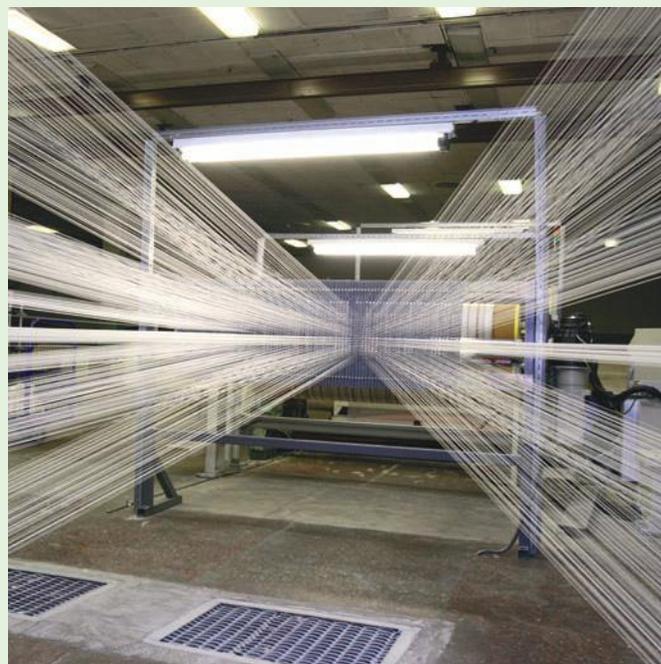
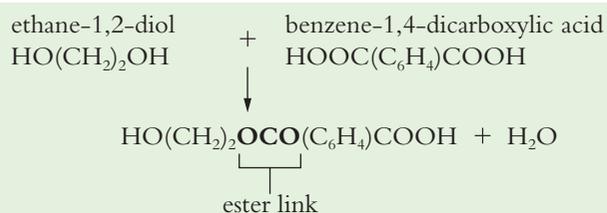
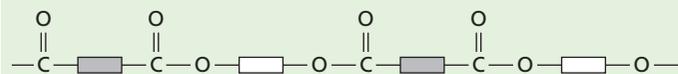


Figure 15.19 Nylon fibre is formed by forcing molten plastic through hundreds of tiny holes.

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.

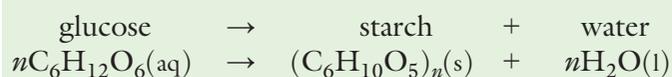
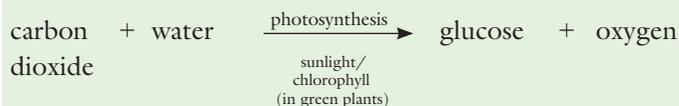
Questions

- Draw the structure of the repeating units found in:
a nylon **b** Terylene.
- 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.



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 $(\text{CH}_2\text{O})_x$.

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).

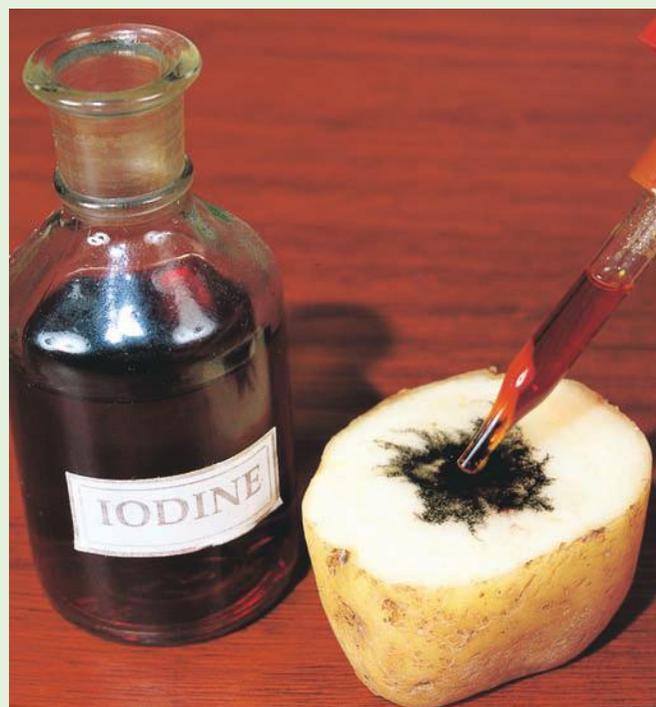
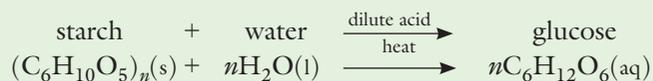


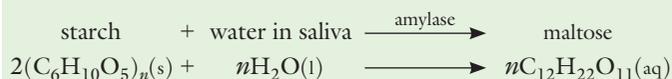
Figure 15.20 A dark blue–black colour is produced when dilute iodine solution is applied to starch, for example in a potato.

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.

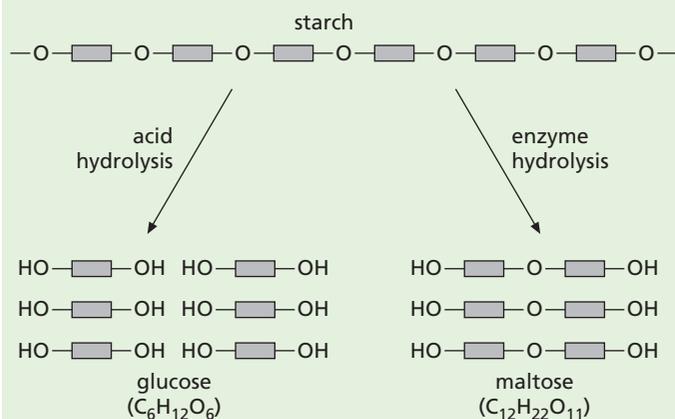


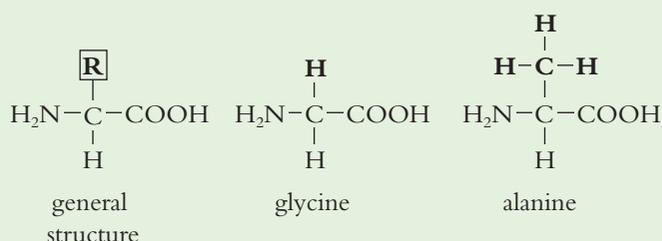
Figure 15.21 Starch produces glucose or maltose depending on the type of hydrolysis used. Using chromatography it is possible to separate and identify the products of the different styles of hydrolysis.

Questions

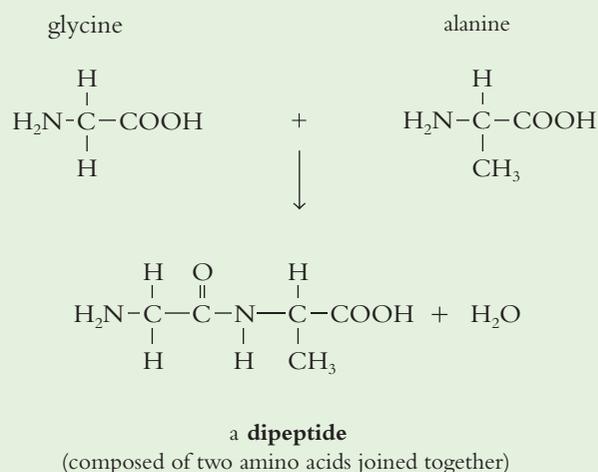
- 1 In the hydrolysis of starch, how, using a chemical test, could you tell whether all the starch had been hydrolysed?
- 2 Describe a method you could possibly use to identify the products of the different types of hydrolysis.

Amino acids

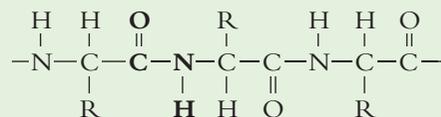
There are 20 different amino acids and they each possess two functional groups. One is the carboxylic acid group, $-\text{COOH}$. The other is the amine group, $-\text{NH}_2$. The two amino acids shown at the top of the next column are glycine and alanine.



Amino acids are the building blocks of **proteins**. Similar to nylon (see p. 241) proteins are polyamides, as they contain the $-\text{CONH}-$ 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.

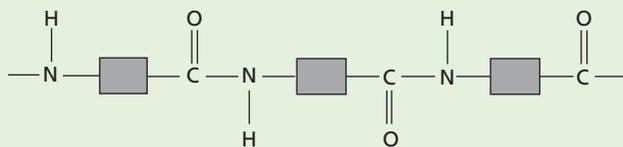


Figure 15.22 General structure of a protein.

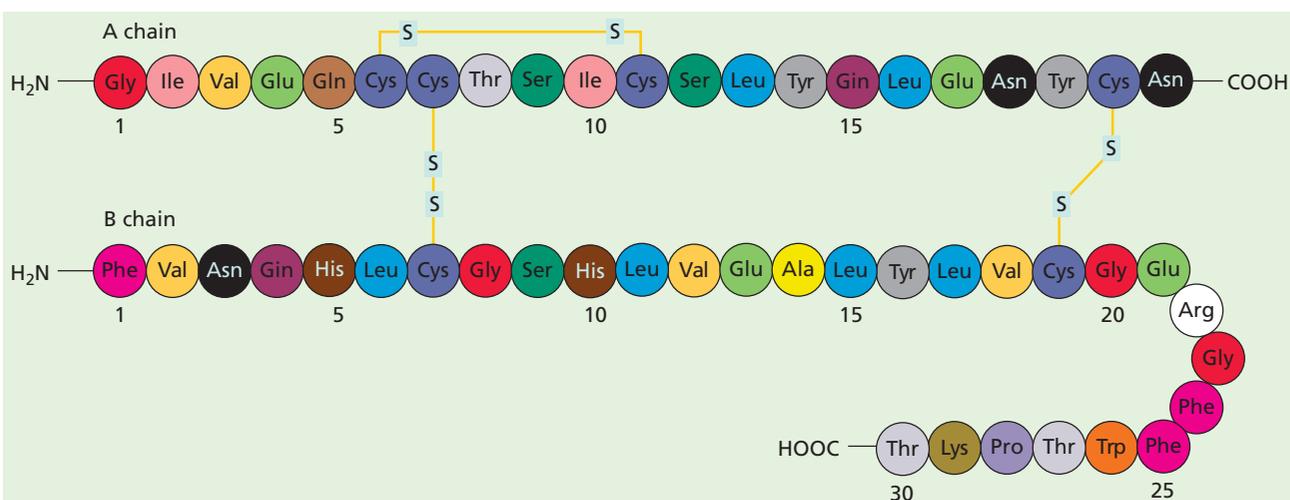


Figure 15.23 The structure of a protein – human insulin (the different coloured circles represent different amino acids in this protein).

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.



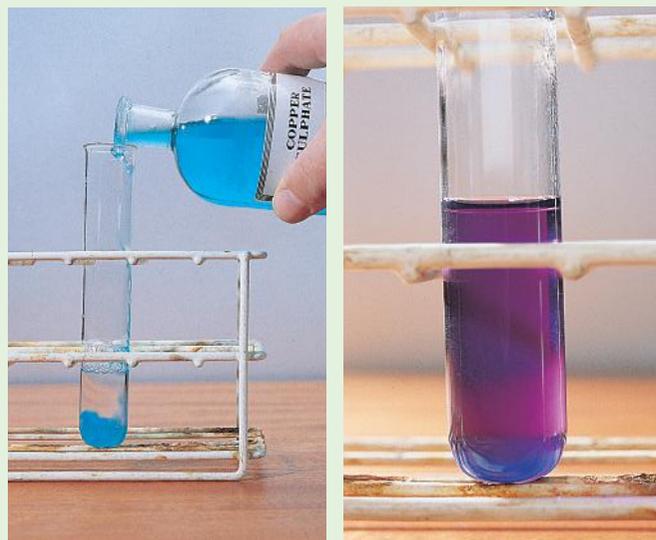
Figure 15.24 Amino acids can be separated and identified by TLC.

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.

DNA

Deoxyribonucleic acid (DNA) belongs to a group of chemicals called the nucleic acids (Figure 15.26). They are also biopolymers. DNA controls the protein 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!



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.

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.

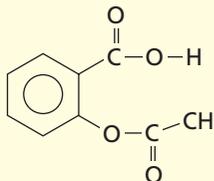
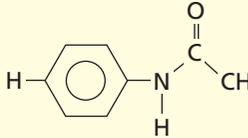
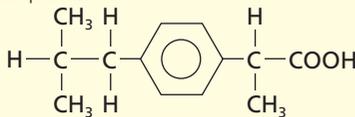
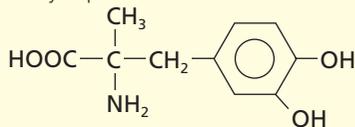
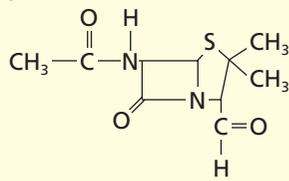
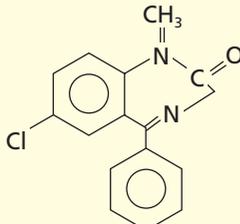
There are, of course, many other types of drug available which have very specific uses. For example, methyldopa was developed to relieve hypertension (high blood pressure), and antihistamines were developed to help control travel sickness, hayfever and allergic reactions.

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)

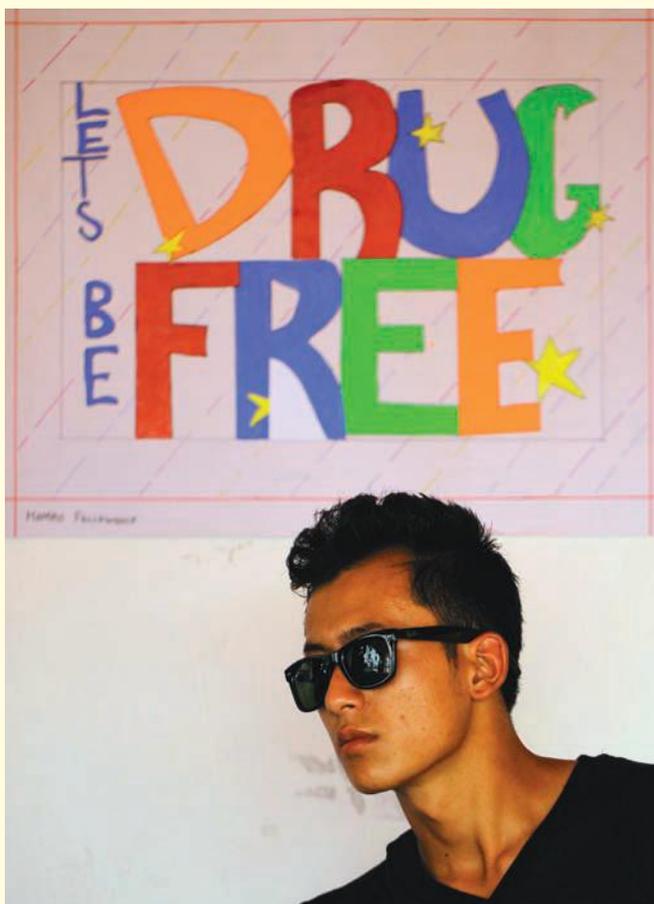


Figure 15.28 Increasing public awareness of the risks associated with drug abuse is very important.

Questions

- Using the data given in Table 15.4, suggest which of the pharmaceuticals contains:

a sulfur	c an -NH_2 group
b an -OH group	d an ester group.
- 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.

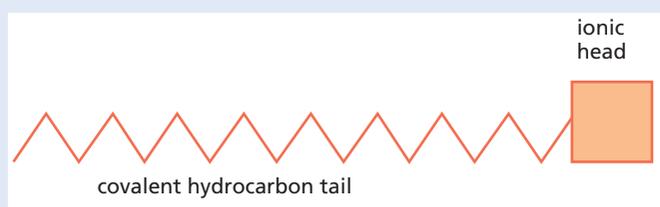
- **Aerobic respiration** Respiration that takes place in the presence of air.
- **Alcohols** Organic compounds containing the -OH group. They have the general formula $\text{C}_n\text{H}_{2n+1}\text{OH}$. Ethanol is by far the most important of the alcohols and is often just called 'alcohol'.
- **Amino 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 $(\text{CH}_2\text{O})_x$.
- **Carboxylic acids** A family of organic compounds containing the functional group -COOH . They have the general formula $\text{C}_n\text{H}_{2n+1}\text{COOH}$. 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, $-\text{NHOC}-$.
- **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

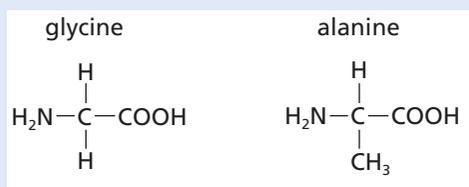
Additional questions

- 1 Explain the following.
- The problem of plastic waste has been overcome.
 - The majority of detergents produced today are biodegradable.
 - In bread making, yeast is added to the mix and the dough left to stand for a period of time.
 - Poly(ethene) is a thermoplastic.
- 2 a A detergent molecule may be represented by the following simplified diagram.

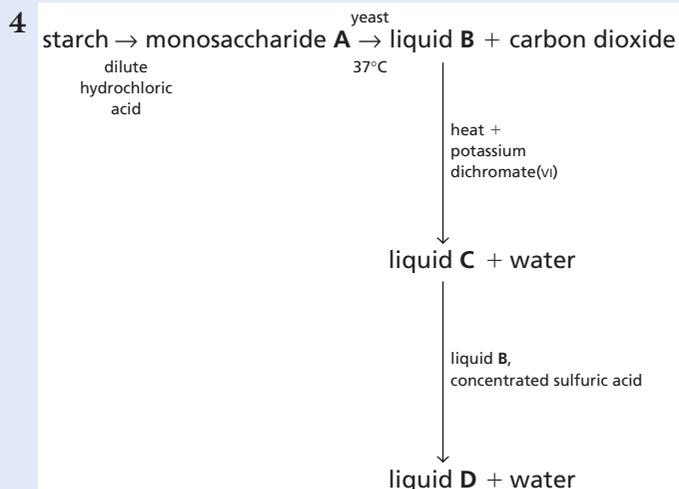


Use this representation of a detergent molecule in a series of labelled diagrams to show how detergents can remove grease from a piece of greasy cloth.

- Explain why detergents do not form a scum with hard water, whereas soaps do.
 - The modern detergents are biodegradable.
 - Explain what this statement means.
 - 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.



- Name the type of polymerisation involved in protein formation.
- Draw a structural formula to represent the part of the protein chain formed by the reaction between the amino acids shown above.
- What is the name given to the common linkage present in protein molecules?
- Why is there such a huge variety of proteins?
- Name and describe the features of the two broad groups of proteins.



- Name and give the formula and one use of each of the substances **A** to **D**.
 - Write word and balanced chemical equations for the reactions involved in the formation of liquids **B**, **C** and **D**.
 - Starch is classified as a natural polymer or 'biopolymer'. Explain the meaning of this statement.
 - Name the processes by which:
 - starch is broken down
 - liquid **B** is formed
 - liquid **C** is formed
 - 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
$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{H} \\ \\ \text{H} \end{array}$		

- Describe a simple chemical test that could be used to distinguish methanol from methanoic acid.
- Name the class of compound produced when methanol reacts with methanoic acid.
 - Name the type of reaction taking place.
 - Write a word and balanced chemical equation for this reaction.
 - Give two uses related to the class of compound formed in this reaction.

- d** The following reaction takes place when methanol is burned:
- $$2\text{CH}_3\text{OH}(\text{l}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$$
- $$\Delta H = -1452 \text{ kJ mol}^{-1}$$
- (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

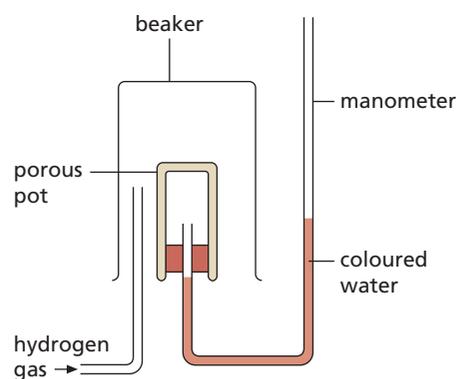
- Measure the melting points of both a pure sample and an impure sample of a solid; this will demonstrate that melting point is a means of assessing purity.
 - Measure the temperature of melting ice, then repeat after adding salt to the ice.
 - Measure the melting point of octadecanoic acid using a water 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

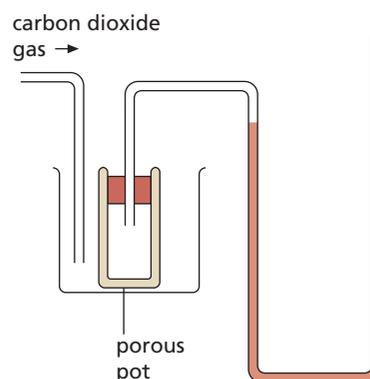
- Students will already be familiar with melting and boiling, so a more profitable approach might be to supply them with data to construct and interpret these curves.

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.

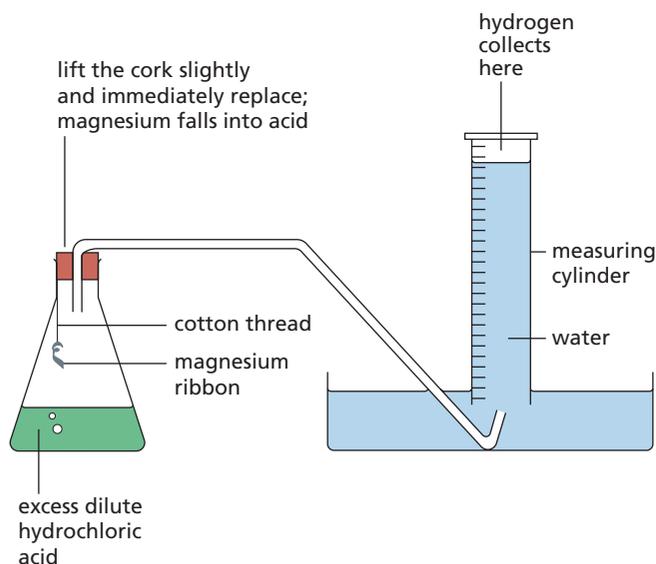


Figure 16.2 Measuring the molar volume of a gas at rtp.

- 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 cm³ of 20 vol hydrogen peroxide and 48 cm³ of water, with 0.2 g of manganese(IV) oxide. Then use a mixture of 4 cm³ of 20 vol hydrogen peroxide and 46 cm³ of water with 0.2 g of manganese dioxide. Finally use a mixture of 2 cm³ of 20 vol hydrogen peroxide and 48 cm³ 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.



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 $\text{MgSO}_4 \cdot x\text{H}_2\text{O}$ or $\text{BaCl}_2 \cdot x\text{H}_2\text{O}$ 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

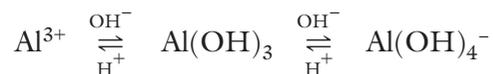
to the aqueous nitrate of another metal. The students look for evidence of displacement and so establish an order of reactivity. Suitable metals are magnesium, zinc, lead, copper and silver. **R**

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.



- 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).

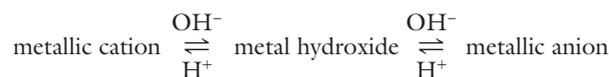
Ammonia gas

- 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

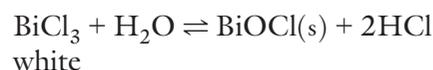
- Acid/base indicators.
- Action of heat on copper(II) sulfate-5-water.
- Add sodium hydroxide solution to a solution of a zinc, an aluminium or a lead salt. The hydroxide will precipitate and then dissolve in excess alkali to form an aluminate, zincate or plumbate. On the

addition of dilute nitric acid the hydroxide will reappear, only to dissolve with the addition of more nitric acid.



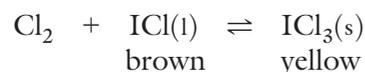
Equilibria

- 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.



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.



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.

- $\text{Fe}^{3+}(\text{aq}) + \text{Ag(s)} \rightleftharpoons \text{Fe}^{2+}(\text{aq}) + \text{Ag}^+(\text{aq})$
- $\text{FeCl}_3 + 3\text{KCNS} \rightleftharpoons \text{Fe(CNS)}_3 + 3\text{KCl}$

Although this equation is not strictly correct, at this level it is preferable to the ionic equation involving FeCNS^{2+} .

- $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$
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.
- $\text{NH}_4\text{Cl} \rightleftharpoons \text{NH}_3 + \text{HCl}$

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 cm³ 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 cm³ 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 – NiCO₃, CoCO₃ and Na₂CO₃(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 cm³ of ethanol and 1 cm³ 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).

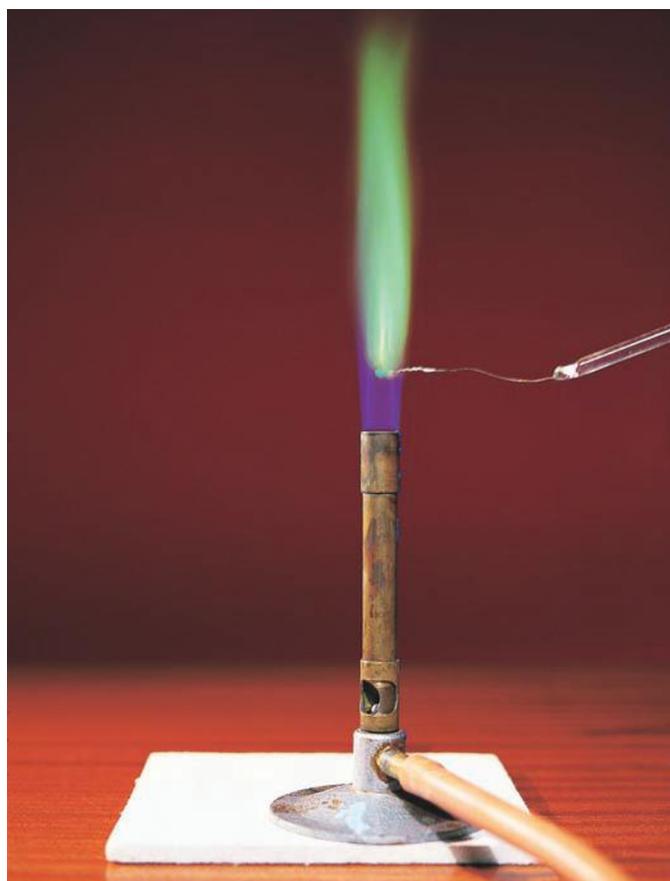


Figure 16.3 The green colour is characteristic of copper.

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	<ul style="list-style-type: none"> • 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	<ul style="list-style-type: none"> • Brown colour • Turns moist indicator paper pink
Water (H ₂ O)	Colourless (odourless)	No effect – neutral	<ul style="list-style-type: none"> • Turns blue cobalt chloride paper pink • Turns anhydrous copper(ii) sulfate from white to blue

Questions to help your understanding

- For each of the following pairs of substances, describe a chemical test you would carry out to distinguish between them.
 - potassium sulfate and potassium sulfite
 - ammonium chloride and aluminium chloride
 - zinc nitrate and calcium nitrate
 - sodium chloride and sodium iodide
 - iron(ii) sulfate and copper(ii) sulfate
- Sodium carbonate hydrate contains water of crystallisation. When it is heated strongly it gives off the water of crystallisation, which can be collected.
 - The substance left behind is anhydrous sodium sulfate. Describe a chemical test to show that this substance contains sodium (cation) and carbonate (anion).
 - Describe two chemical tests to show that the colourless liquid produced and collected is indeed water.
 - 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, 20cm^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 0.0cm^3 mark. 5cm^3 of acid **C** was added to the sodium hydroxide in the flask. The temperature of the mixture was measured.

Further 5cm^3 portions of acid **C** were added to the mixture in the flask, stirring with the thermometer until a total volume of 30cm^3 of acid **C** had been added. The temperatures after each 5cm^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/ $^{\circ}\text{C}$
0		
5		
10		
15		
20		

Volume of acid C added/ cm^3	Thermometer diagrams	Temperature/ $^{\circ}\text{C}$
25		
30		

[2]

Experiment 2

The burette was emptied and rinsed with water. Experiment 1 was repeated using acid **D**.

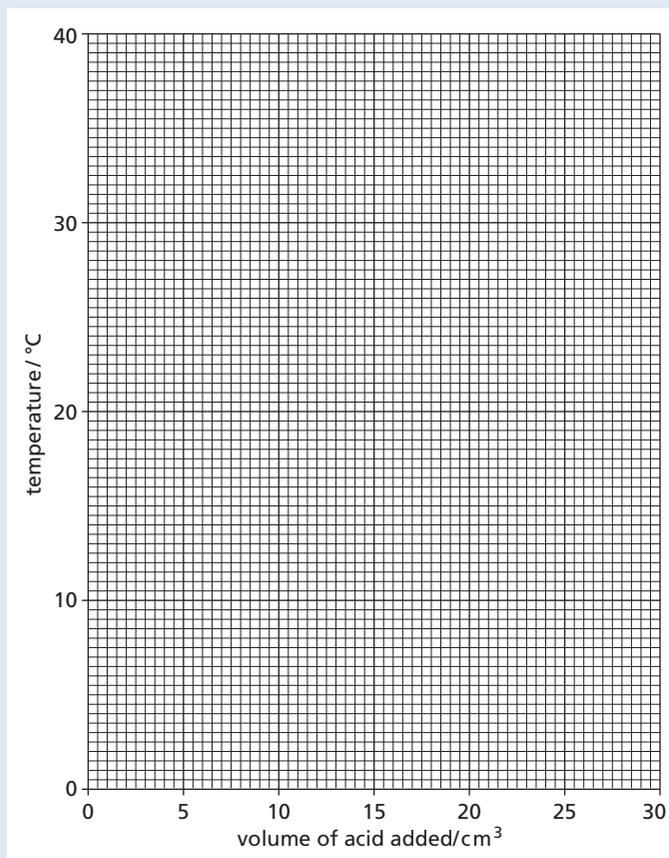
b Use the thermometer diagrams to record the temperatures in the table of results.

Table of results

Volume of acid D added/ cm^3	Thermometer diagrams	Temperature/ $^{\circ}\text{C}$
0		
5		
10		
15		
20		
25		
30		

[2]

- c Plot the results for Experiments 1 and 2 on the grid and draw two smooth line graphs. Clearly label your graphs.



- 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. [6]
- 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.

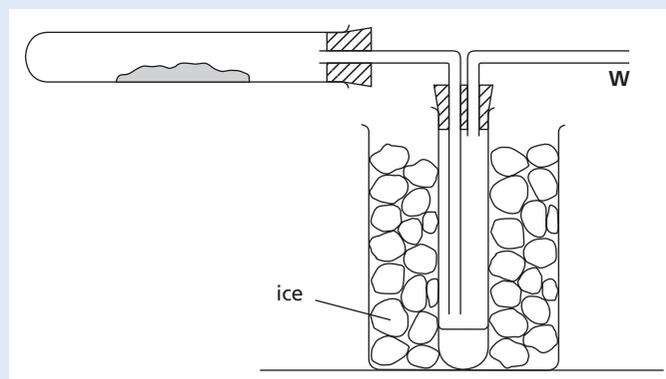
Tests	Observations
Tests on solid E	
a Appearance of solid E	white crystalline solid
b A little of solid E was heated in a test-tube.	colourless drops of liquid formed at the top of the tube
c A little of solid E was dissolved in distilled water. The solution was divided into four test-tubes and the 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, $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$, 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)

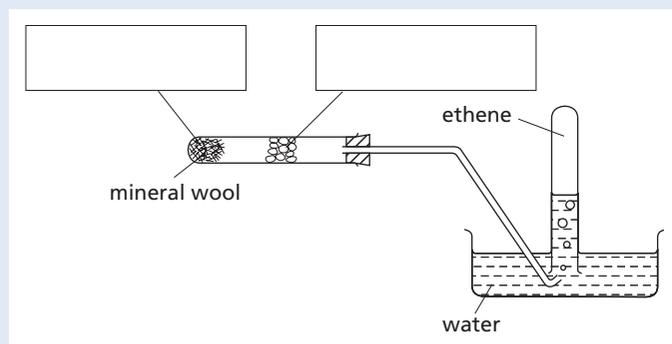
- 4 Malachite is a naturally occurring form of copper carbonate. Outline how a sample of copper metal could be obtained from large lumps of malachite in the laboratory.

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, $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$.

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.

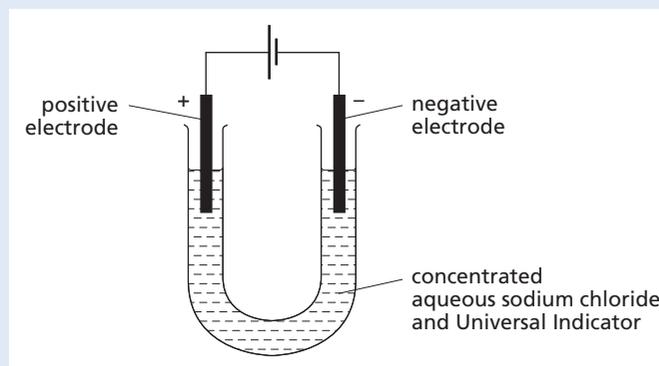
Step 4 Heat the solution until crystals form. Obtain the crystals and dry them.

- 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. A glowing splint was inserted into the tube.	slow effervescence splint relit
c Dilute sulfuric acid was added to solid S in a test-tube. The mixture was heated to boiling point. The solution was divided into three equal portions into test-tubes.	blue solution formed
(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. Excess ammonia solution was then added. [2] [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. A glowing splint was inserted into the tube.	rapid effervescence 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.

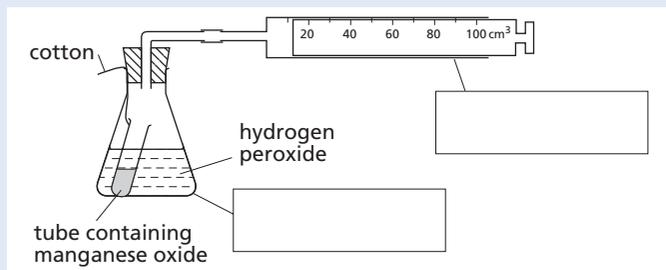
a Plan an experiment to obtain an aqueous solution of an indicator from some berries. [3]

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)

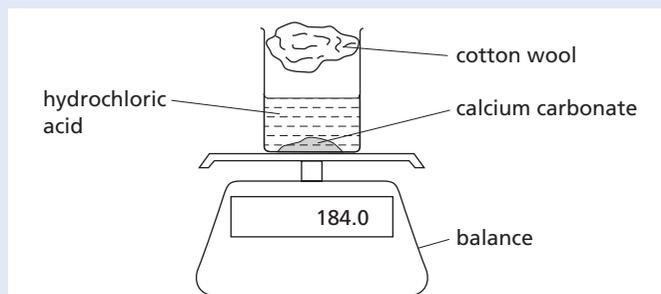
13 Three unlabelled bottles of chemicals each contained one of the following liquids:

- 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.

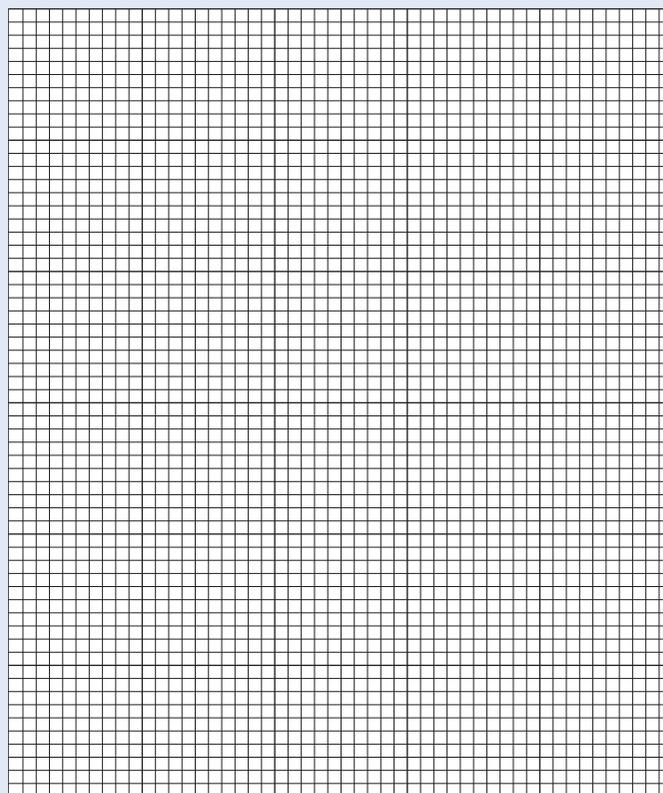


The beaker was placed on a balance and the mass of the beaker and contents recorded every minute.

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)

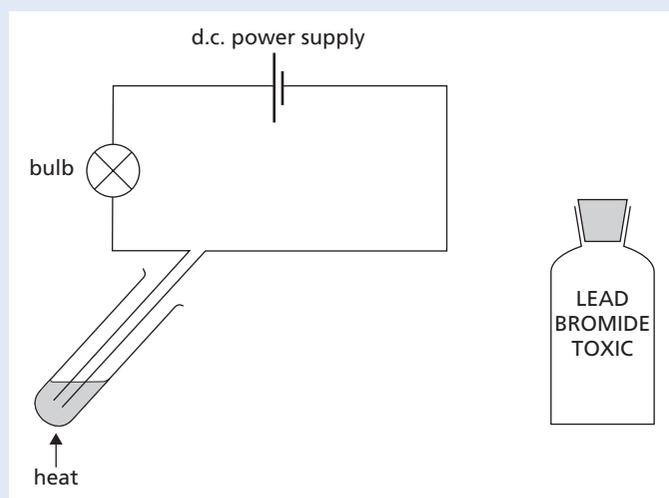
- 15 Leaves from trees contain a mixture of coloured pigments which are not soluble in water. A student was given these two instructions to investigate the pigments in the leaves.

- 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.

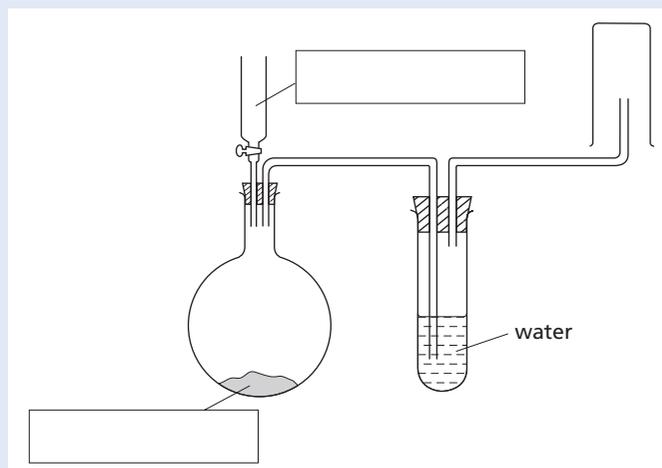


- a (i) Clearly label the electrodes on the diagram. [1]
 (ii) Suggest a suitable material to make the electrodes. [1]
 b Give two observations expected when the lead bromide is heated to melting point. [2]
 c State two different safety precautions when carrying out this experiment. [2]

[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.

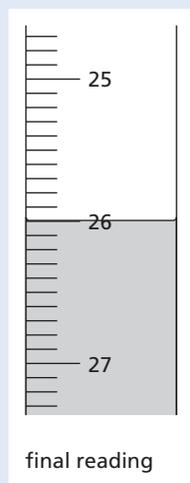
Two experiments were carried out.

Experiment 1

- a About 1 cm³ of aqueous sodium hydroxide was added to a little of the salt solution **A** and the observation noted.

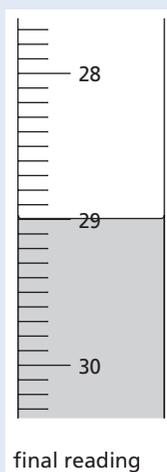
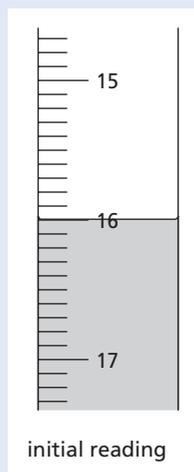
observation *green precipitate formed*

- b** A burette was filled with potassium manganate(VII) solution up to the 0.0 cm³ mark. By using a measuring cylinder, 25 cm³ 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 cm³ 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/cm³

	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 cm³ 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

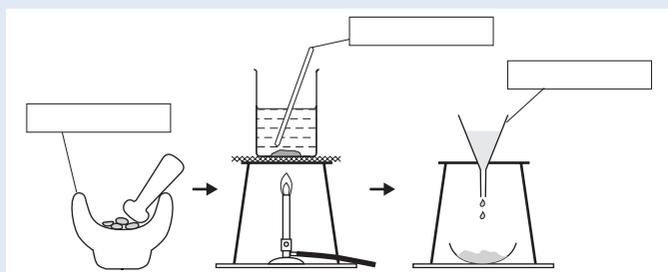
- a** Give a chemical test to show the presence of ammonia in Bite Relief solution.
 test _____
 result _____ [2]
- b** What practical method could be used to separate the mixture of alcohol (bp 78 °C) and water (bp 100 °C)? [2]

- 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]

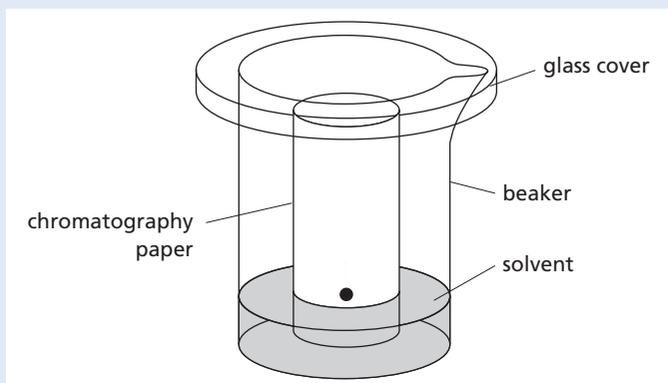
(Cambridge IGCSE Chemistry 0620 Paper 06
Q7 June 2008)

- 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.

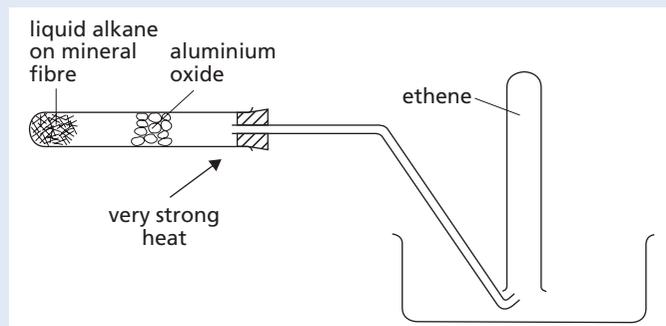


- b (i) Name the solvent used. [1]
(ii) Label, with an arrow, the origin on the diagram. [1]
- c Sketch the chromatogram you would expect if two different colours were present in the sweets. [1]

[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]

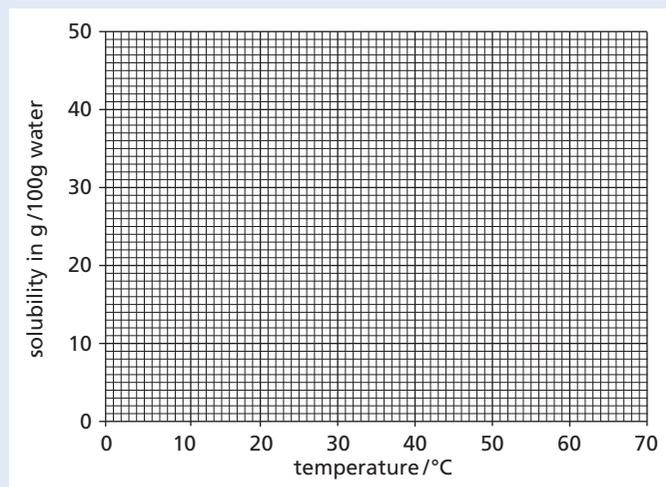
[Total: 4]

(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/100g 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)

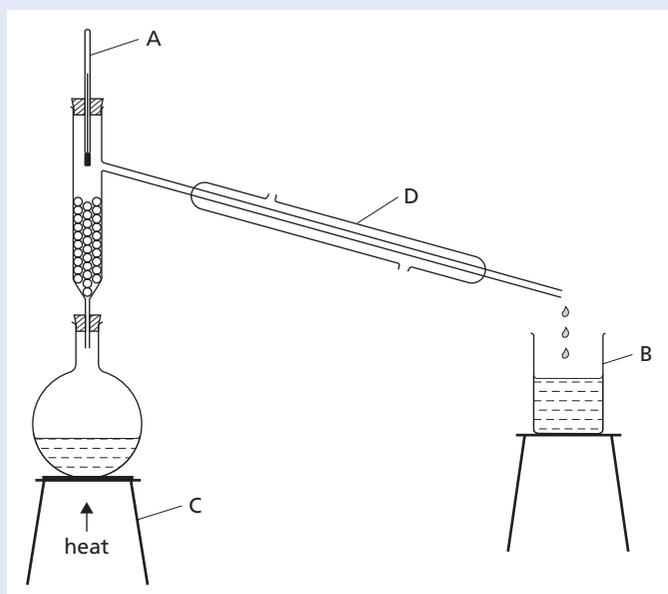
23 A solution of magnesium sulfate can be made by reacting magnesium oxide with warm sulfuric acid.

- 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, $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$, 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]

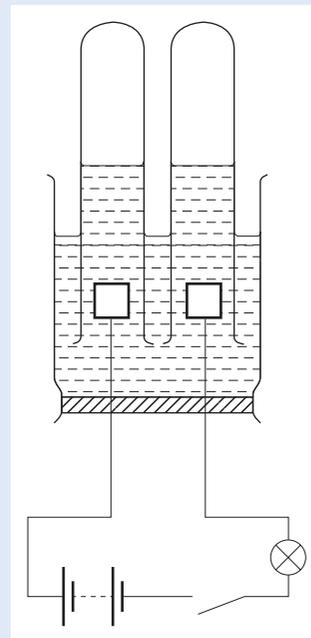
- b** What is the purpose of **D**? [1]
- c** How could the purity of the ethanol collected be checked? [1]

[Total: 5]

(Cambridge IGCSE Chemistry 0620 Paper 06

Q1 June 2007)

25 Concentrated hydrochloric acid can be electrolysed using the apparatus shown.



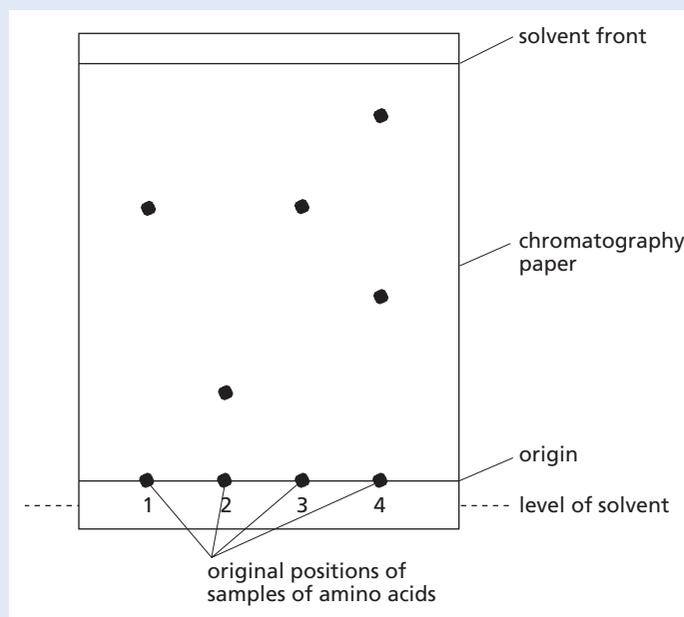
- a** Label the position of the electrodes on the diagram. [1]
- b** Give two observations when the circuit is switched on. [2]
- c** (i) Name the product at the positive electrode. [1]
- (ii) State a test for this product and the result of the test.
test _____
result _____ [2]

[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.

($1 \text{ dm}^3 = 1000 \text{ cm}^3$)

[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 pH5 R pH7
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]

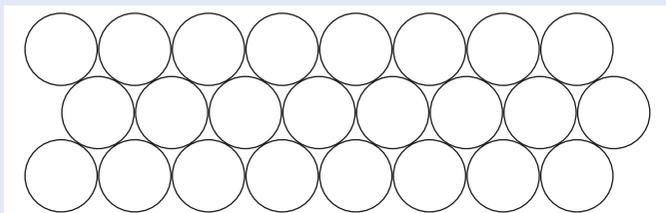
(Cambridge IGCSE Chemistry 0620 Paper 21
Q1 November 2010)

2 About 4000 years ago the Bronze Age started in Britain. Bronze is an alloy of copper and tin.

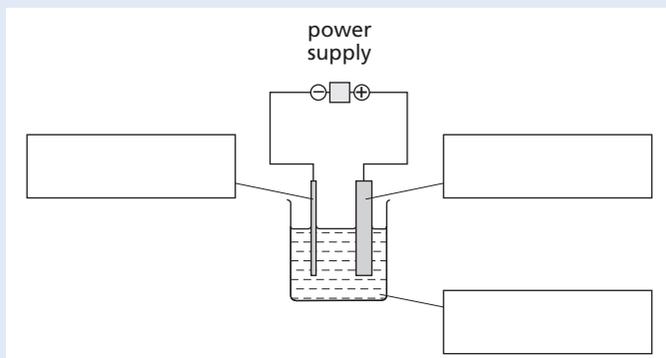
a (i) Suggest a reason why a bronze axe was better than a copper axe. [1]

(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, $\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2$.
- (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.

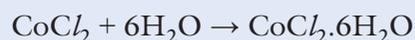
Step 1 Add an excess of cobalt(II) carbonate to hot dilute hydrochloric acid.

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.



Maximum yield

Number of moles of HCl used = _____

Number of moles of CoCl_2 formed = _____

Number of moles of $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ formed = _____

Mass of one mole of $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ = 238 g

Maximum yield of $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ = _____ 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

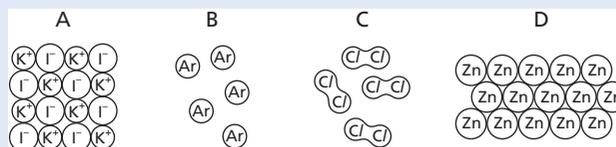
giant ionic

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.

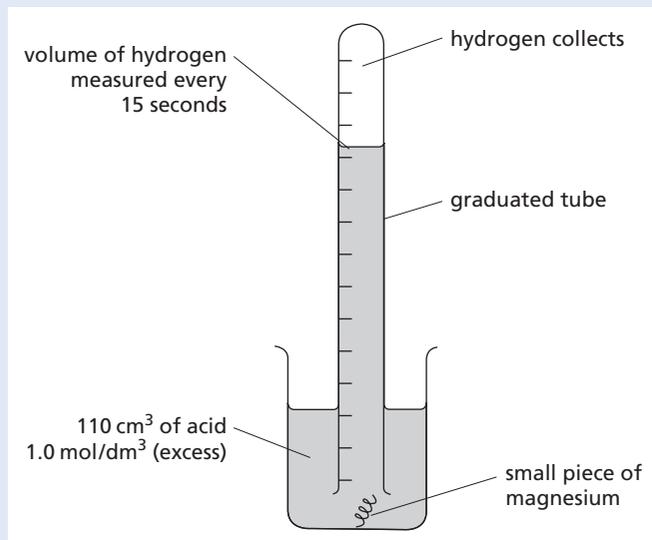
electrons ionic molecular molten solid

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]

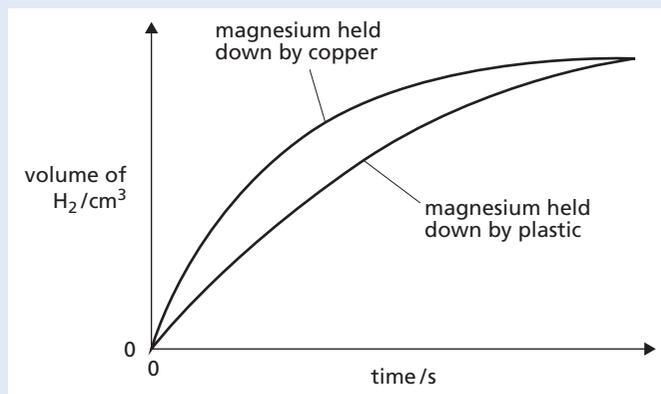
(Cambridge IGCSE Chemistry 0620 Paper 22
 Q2 June 2010)

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.

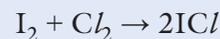


- (i) In which experiment did the magnesium react faster? [1]
- (ii) Suggest a reason why the experiment chosen in (i) had the faster rate. [1]
- c** The experiment was repeated using 1.0 mol/dm³ propanoic acid instead of 1.0 mol/dm³ hydrochloric acid. Propanoic acid is a weak acid.
- (i) How would the 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]

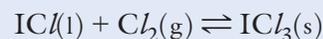
[Total: 10]

(Cambridge IGCSE Chemistry 0620 Paper 31
 Q3 June 2010)

6 Iodine reacts with chlorine to form dark brown iodine monochloride.



This reacts with more chlorine to give yellow iodine trichloride. There is an equilibrium between these iodine chlorides.



- 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]

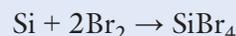
(Cambridge IGCSE Chemistry 0620 Paper 31
Q3 June 2009)

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.



(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]

atom	a substance containing different atoms or ions bonded together
compound	a substance made up of one type of atom
element	the smallest part of an element which takes part in a chemical reaction
ion	the smallest group of covalently bonded atoms which can exist on its own
molecule	a charged atom or group of atoms

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

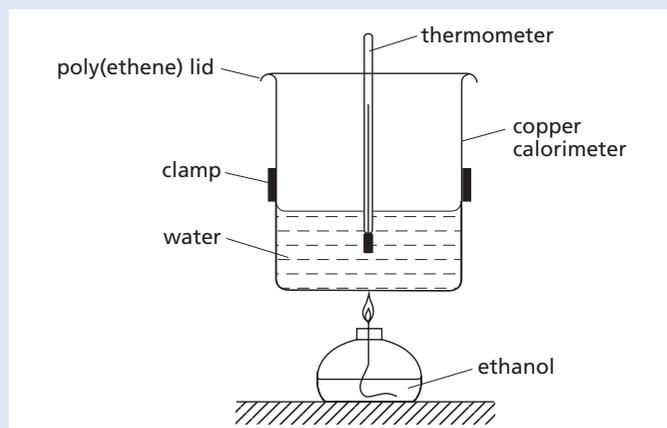
helium has an incomplete outer shell of electrons

[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 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.



(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]

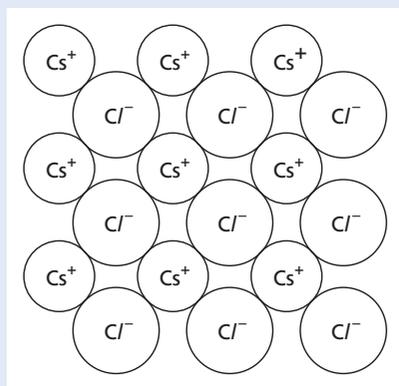
(Cambridge IGCSE Chemistry 0620 Paper 02
Q3 November 2009)

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/cm ³	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		

- e The diagram shows the structure of caesium chloride. [2]



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/dm ³
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 SO₄²⁻. [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 XH₄.
 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

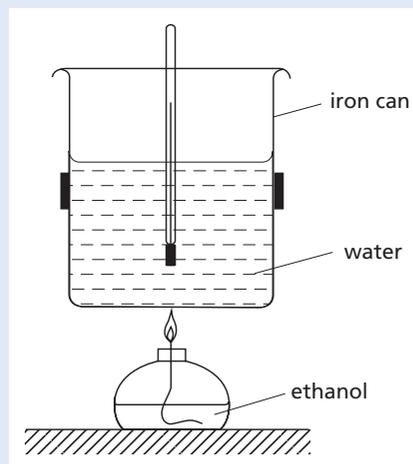
From this list, choose the most suitable technique to separate the following.

- 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]

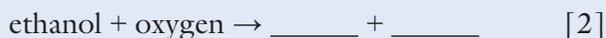
[Total: 6]

(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.



- 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]

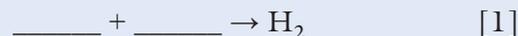
[Total: 14]

(Cambridge IGCSE Chemistry 0620 Paper 31
 Q2 November 2008)

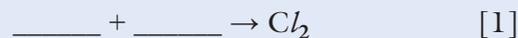
20 The electrolysis of concentrated aqueous sodium chloride produces three commercially important chemicals: hydrogen, chlorine and sodium hydroxide.

- a The ions present are $\text{Na}^+(\text{aq})$, $\text{H}^+(\text{aq})$, $\text{Cl}^-(\text{aq})$ and $\text{OH}^-(\text{aq})$.

- (i) Complete the ionic equation for the reaction at the negative electrode (cathode).



- (ii) Complete the ionic equation for the reaction at the positive electrode (anode).



- (iii) Explain why the solution changes from sodium chloride to sodium hydroxide. [1]

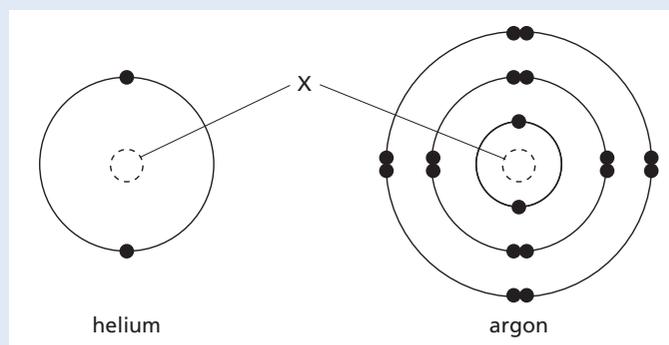
- b (i) Why does the water supply industry use chlorine? [1]
 (ii) Name an important chemical that is made from hydrogen. [1]
 (iii) How is sodium hydroxide used to make soap? [2]

[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.

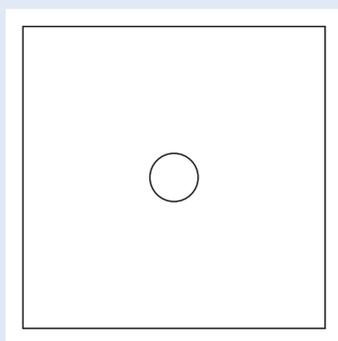
Helium has an incomplete outer shell of electrons.

(iii) How many protons are there in an atom of argon? [1]

(iv) The symbol for a particular isotope of helium is written as ${}^4_2\text{He}$. [1]

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.



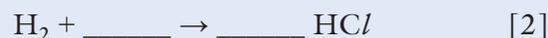
[2]

[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.



b Draw a dot and cross diagram for a molecule of hydrogen chloride. Show all the electrons.

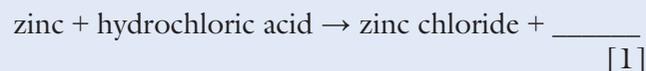
Use **o** for an electron from a hydrogen atom.

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.



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) $\text{Mg} + \text{H}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{H}_2$
Change into a word equation. [1]

(ii) lithium oxide + sulfuric acid \rightarrow lithium sulfate + water
Change into a symbol equation. [2]

(iii) $\text{CuO} + 2\text{H}^+ \rightarrow \text{Cu}^{2+} + \text{H}_2\text{O}$
Change the ionic equation into a symbol equation. [2]

(iv) $\text{Na}_2\text{CO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + \text{CO}_2 + \text{H}_2\text{O}$
Change into a word equation. [1]

b When sulfuric acid dissolves in water, the following reaction occurs.



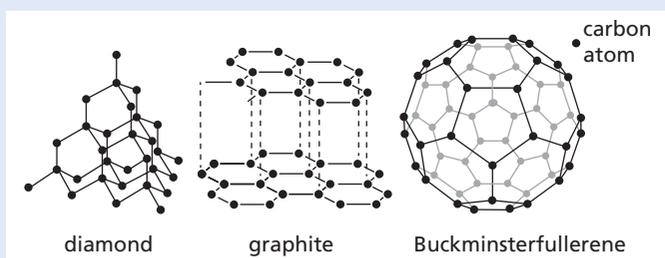
Explain why water is behaving as a base in this reaction. [2]

- 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, CH₄.
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]

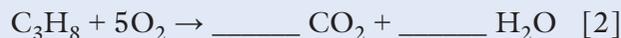
(Cambridge IGCSE Chemistry 0620 Paper 22
Q7 June 2010)

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:



- (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).

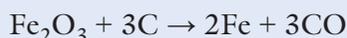
calcium carbonate carbon dioxide
 hydrogen chloride
 iron(III) oxide lead(II) bromide
 methane sodium hydroxide

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.



- (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]

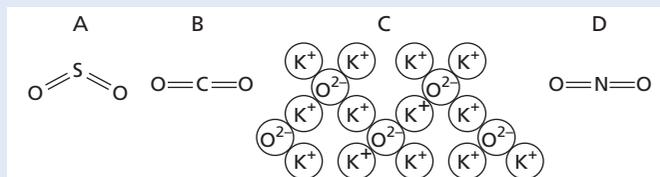
(Cambridge IGCSE Chemistry 0620 Paper 02
 Q1 June 2009)

- 27 The list shows some non-metallic elements.

bromine
 carbon
 fluorine
 krypton
 nitrogen
 oxygen

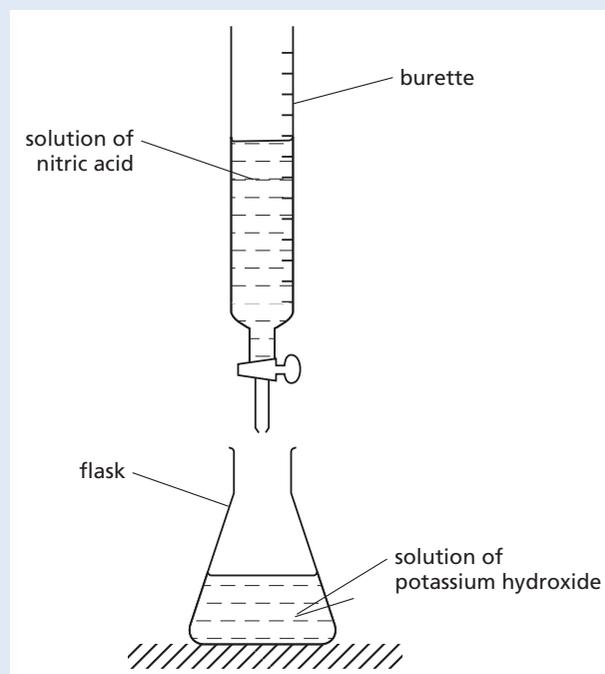
- a Which **two** elements in the list are in the same Group of the Periodic Table? [1]
 b Which element in the list has the highest proton number? [1]

- 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.

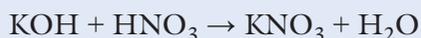
A student used the apparatus shown to add an aqueous solution of nitric acid to an aqueous solution of potassium hydroxide. He added the acid until it was in excess.



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



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

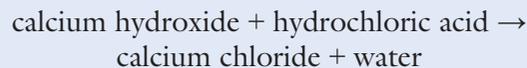
[1]

c Some farmers use calcium hydroxide to control soil acidity.

(i) Why is it important to control soil acidity? [1]

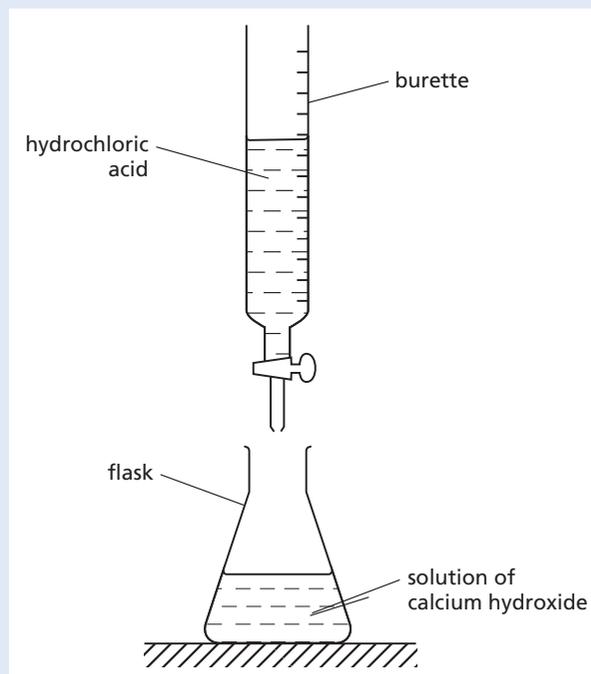
(ii) Acid rain can cause soil to become acidic.
Describe how acid rain is formed. [3]

d Calcium hydroxide reacts with hydrochloric acid.



(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]

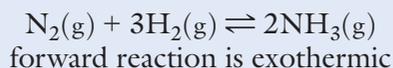
(Cambridge IGCSE Chemistry 0620 Paper 21
Q5 November 2010)

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.



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.



- (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]

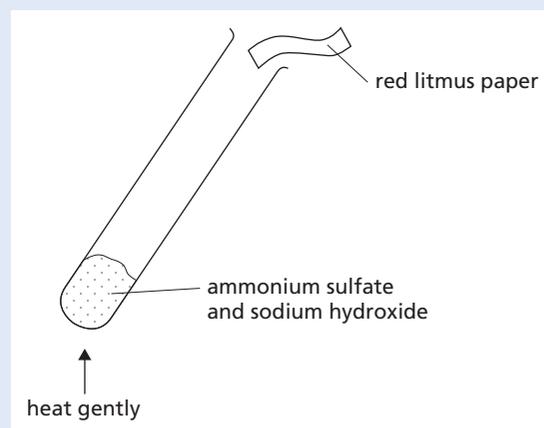
(Cambridge IGCSE Chemistry 0620 Paper 22
 Q6 June 2010)

- 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 + hydrochloric → ___ + ___ + ___
 carbonate 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.

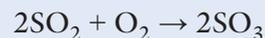
Type of oxide	pH of solution oxide	Example
acidic		
basic		
neutral		

- 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)

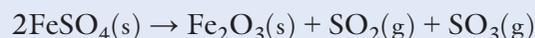
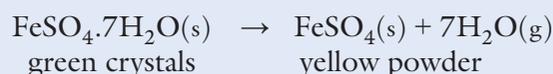
- 35 a Sulfuric acid is made by the Contact process.



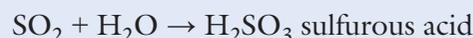
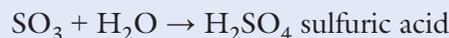
This is carried out in the presence of a catalyst at 450 °C and 2 atmospheres pressure.

- (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 green vitriol, $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$. The gases formed were cooled.

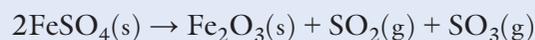


On cooling



- (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.



mass of one mole of $\text{FeSO}_4 = 152 \text{ g}$

number of moles of FeSO_4 used = _____

number of moles of Fe_2O_3 formed = _____

mass of one mole of $\text{Fe}_2\text{O}_3 = \text{_____ g}$

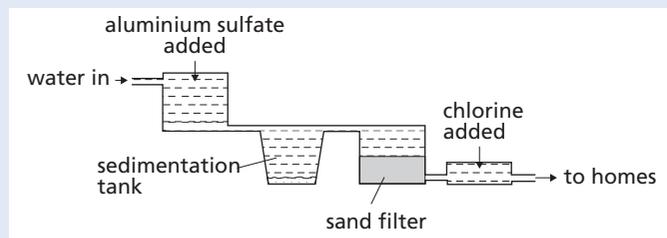
mass of iron(III) oxide formed = _____ g
 number of moles of SO_3 formed = _____
 volume of sulfur trioxide formed = _____ dm^3

[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.



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.



(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 $Fe_2O_3 \cdot 2H_2O$. 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 $Fe_2O_3 \cdot 2H_2O$. [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.



[Total: 11]

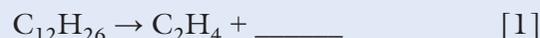
(Cambridge IGCSE Chemistry 0620 Paper 31
Q2 June 2010)

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, $C_{12}H_{26}$.



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)

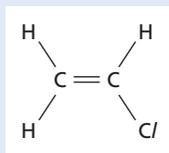
42 Monomers polymerise to form polymers or macromolecules.

a (i) Explain the term *polymerise*. [1]

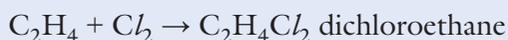
(ii) There are two types of polymerisation – addition and condensation.

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.



This is heated to make chloroethene.



(i) Ethene is made by cracking alkanes. Complete the equation for cracking dodecane.



Another method of making dichloroethane is from ethane.



(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)

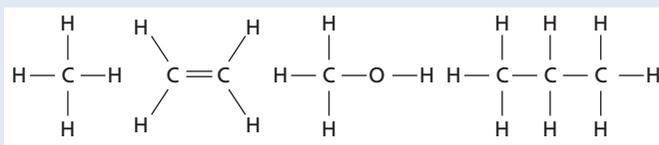
43 Petroleum is a mixture of hydrocarbons which can be separated into fractions such as petrol, paraffin and diesel.

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,

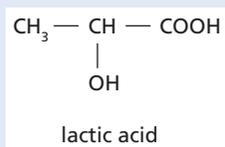
(ii) hydrocarbon? [2]

[Total: 11]

(Cambridge IGCSE Chemistry 0620 Paper 02

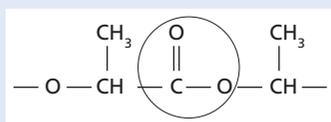
Q3 June 2009)

44 Lactic acid can be made from corn starch.

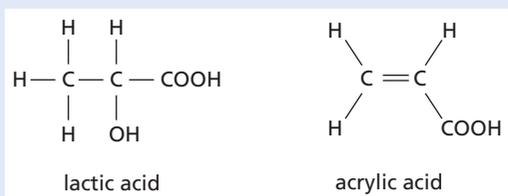


It polymerises to form the polymer, polylactic acid (PLA) which is biodegradable.

- a Suggest **two** advantages that PLA has compared with a polymer made from petroleum.
- 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 \rightarrow _____ + _____ [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)

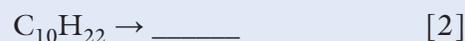
- [2] 45 Butan-1-ol is used as a solvent for paints and varnishes, to make esters and as a fuel. Butan-1-ol can be manufactured from but-1-ene, which is made from petroleum.

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, $\text{C}_{10}\text{H}_{22}$, by cracking.

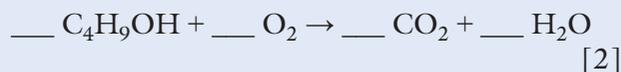
(i) Give the reaction conditions. [2]

(ii) Complete an equation for the cracking of decane, $\text{C}_{10}\text{H}_{22}$, to give but-1-ene.



(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.



(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]

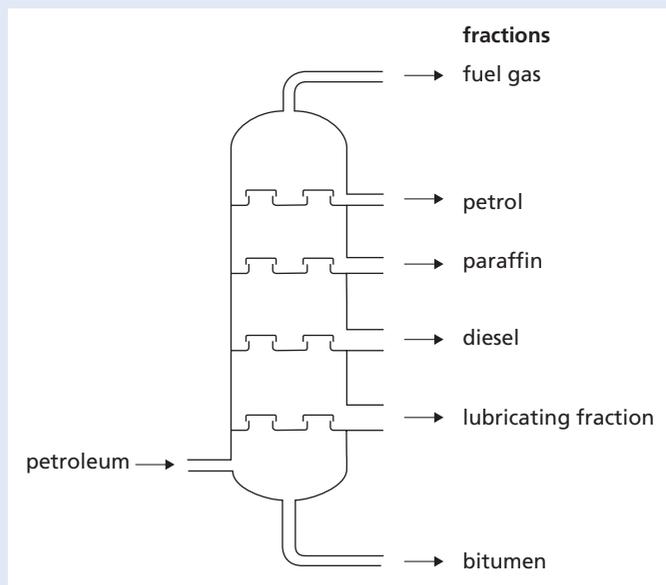
(ii) Why is it important to develop these fuels, such as biobutanol, as alternatives to petroleum? [1]

- 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)

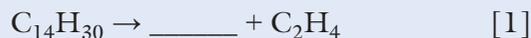
- 46 Petroleum is separated into useful fractions by distillation.



- 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 Ethanol is formed when steam reacts with ethene at high pressure and temperature. A catalyst of phosphoric acid is used.



- (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]

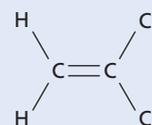
- (iv) Phosphoric acid is a typical acid. State what you would observe when a solution of phosphoric acid is added to blue litmus, a solution of sodium carbonate. [2]

[Total: 13]

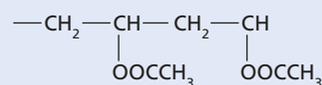
(Cambridge IGCSE Chemistry 0620 Paper 02
 Q6 June 2008)

- 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.



- 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 cm³ of butane is mixed with 100 cm³ 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?

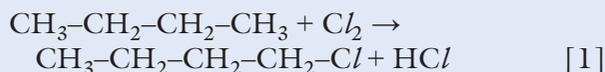


Volume of oxygen left = _____ cm³

Volume of carbon dioxide formed = _____ cm³ [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.



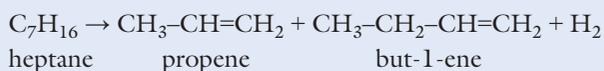
(i) Name the organic product. [1]

(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.



(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 elements

Group																	
I	II											III	IV	V	VI	VII	0
<div style="display: flex; justify-content: space-between; align-items: center;"> <div style="border: 1px solid black; padding: 5px; text-align: center;"> 1 H Hydrogen 1 </div> <div style="border: 1px solid black; padding: 5px; text-align: center;"> 4 He Helium 2 </div> </div>																	
7 Li Lithium 3	9 Be Beryllium 4											11 B Boron 5	12 C Carbon 6	14 N Nitrogen 7	16 O Oxygen 8	19 F Fluorine 9	20 Ne Neon 10
23 Na Sodium 11	24 Mg Magnesium 12											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
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	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
85 Rb Rubidium 37	88 Sr Strontium 38	89 Y Yttrium 39	91 Zr Zirconium 40	93 Nb Niobium 41	96 Mo Molybdenum 42	96 Tc Technetium 43	101 Ru Ruthenium 44	103 Rh Rhodium 45	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
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 Iridium 77	195 Pt Platinum 78	197 Au Gold 79	201 Hg Mercury 80	204 Tl Thallium 81	207 Pb Lead 82	209 Bi Bismuth 83	209 Po Polonium 84	209 At Astatine 85	209 Rn Radon 86
87 Fr Francium 87	226 Ra Radium 88	227 Ac Actinium 89 †															

*58-71 Lanthanoid series

†90-103 Actinoid series

Key

a	a = relative atomic mass
X	X = atomic symbol
b	b = proton (atomic) number

140 Ce Cerium 58	141 Pr Praseodymium 59	144 Nd Neodymium 60	Pm Promethium 61	150 Sm Samarium 62	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
232 Th Thorium 90	Pa Protactinium 91	238 U Uranium 92	Np Neptunium 93	Pu Plutonium 94	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 Lawrencium 103

The volume of one mole of any gas is 24 dm³ at room temperature and pressure (r.t.p.).

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