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Chemistry

Revision Guide

Roger Norris

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How to use this book: a guided tour

Learning outcomes

By the end of this unit you should:

- Understand the importance of purity in substances in everyday life
- Assess the purity of substances from melting and boiling point data
- Describe paper chromatography and interpret simple chromatograms
- Interpret simple chromatograms, including the use of R_f values
- Understand how locating agents are used to make colourless substances visible on a chromatogram
- Describe and explain how solvents, filtration, crystallisation and distillation are used to separate or purify substances
- Suggest suitable purification techniques to obtain a given product

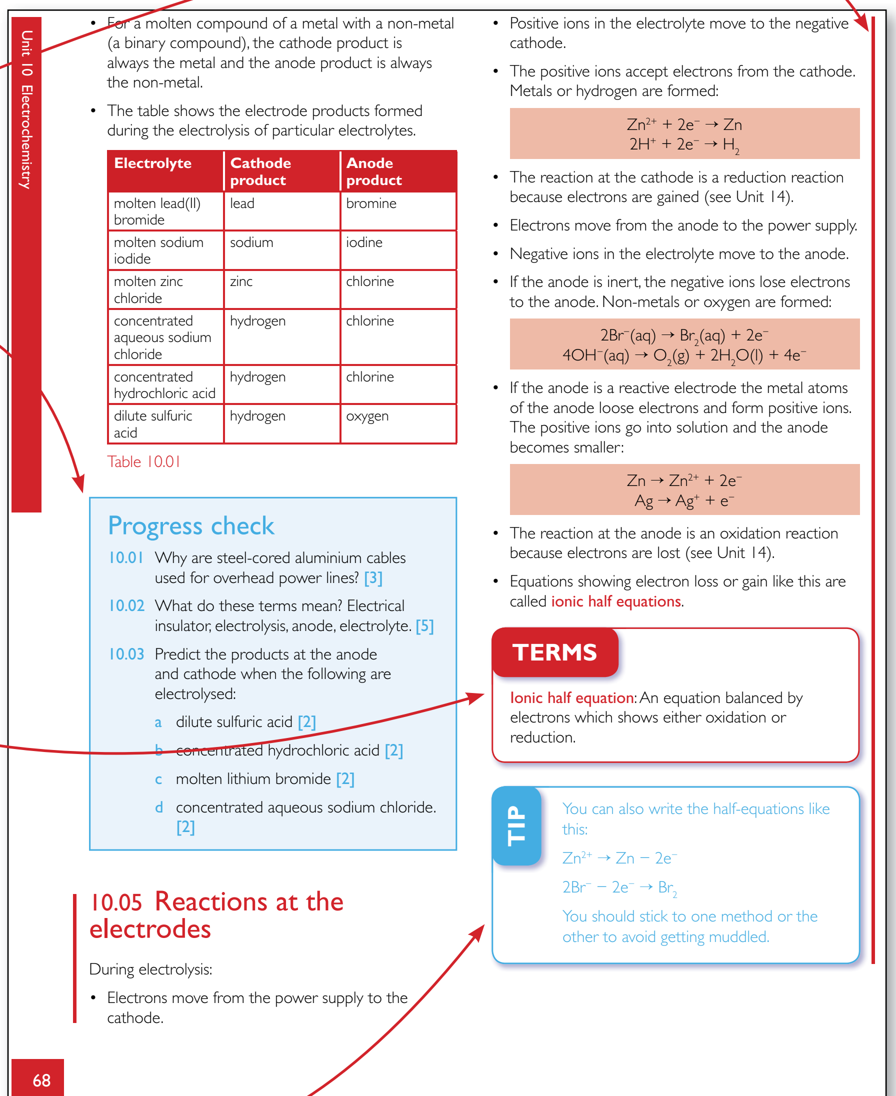
Learning outcomes – set the scene of each chapter, help with navigation through the book and give a reminder of what is important about each topic

Supplement material – indicated by a bold vertical line. This is for students who are taking the Extended syllabus covering the Core and Supplement content

Progress check – check your own knowledge and see how well you are getting on by answering regular questions

Terms – clear and straightforward explanations are provided for the most important words in each topic

Tip – quick suggestions to remind you about key facts and highlight important points



Unit 10 Electrochemistry

- For a molten compound of a metal with a non-metal (a binary compound), the cathode product is always the metal and the anode product is always the non-metal.
- The table shows the electrode products formed during the electrolysis of particular electrolytes.

Electrolyte	Cathode product	Anode product
molten lead(II) bromide	lead	bromine
molten sodium iodide	sodium	iodine
molten zinc chloride	zinc	chlorine
concentrated aqueous sodium chloride	hydrogen	chlorine
concentrated hydrochloric acid	hydrogen	chlorine
dilute sulfuric acid	hydrogen	oxygen

Table 10.01

Progress check

10.01 Why are steel-cored aluminium cables used for overhead power lines? [3]

10.02 What do these terms mean? Electrical insulator, electrolysis, anode, electrolyte. [5]

10.03 Predict the products at the anode and cathode when the following are electrolysed:

a dilute sulfuric acid [2]

b concentrated hydrochloric acid [2]

c molten lithium bromide [2]

d concentrated aqueous sodium chloride. [2]

10.05 Reactions at the electrodes

During electrolysis:

- Electrons move from the power supply to the cathode.
- Positive ions in the electrolyte move to the negative cathode.
- The positive ions accept electrons from the cathode. Metals or hydrogen are formed:

$$Zn^{2+} + 2e^- \rightarrow Zn$$
$$2H^+ + 2e^- \rightarrow H_2$$

- The reaction at the cathode is a reduction reaction because electrons are gained (see Unit 14).
- Electrons move from the anode to the power supply.
- Negative ions in the electrolyte move to the anode.
- If the anode is inert, the negative ions lose electrons to the anode. Non-metals or oxygen are formed:

$$2Br^-(aq) \rightarrow Br_2(aq) + 2e^-$$
$$4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$$

- If the anode is a reactive electrode the metal atoms of the anode loose electrons and form positive ions. The positive ions go into solution and the anode becomes smaller:

$$Zn \rightarrow Zn^{2+} + 2e^-$$
$$Ag \rightarrow Ag^+ + e^-$$

- The reaction at the anode is an oxidation reaction because electrons are lost (see Unit 14).
- Equations showing electron loss or gain like this are called ionic half equations.

TERMS

Ionic half equation: An equation balanced by electrons which shows either oxidation or reduction.

TIP

You can also write the half-equations like this:

$Zn^{2+} \rightarrow Zn - 2e^-$

$2Br^- - 2e^- \rightarrow Br_2$

You should stick to one method or the other to avoid getting muddled.

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Worked example 15.01

a What could you do to tell the difference between an acidic and alkaline solution? [2]

The best ways involve chemistry not taste! Either the litmus test or the pH test is suitable. Do not suggest indicators other than litmus or universal indicator because if the acid or base is very weak, a colour change may not be seen.

Acids turn blue litmus red [1]; alkalis turn red litmus blue [1].

b Name three different compounds you could add to hydrochloric acid to make the salt calcium chloride. Write word equations for the formation of the salt. [4]

Since it is calcium chloride you are making, the substances must be calcium compounds. But not calcium metal because that is not a compound. When writing the equations, concentrate on the products. Do not forget that with a carbonate, salt + carbon dioxide + water are formed. And do not forget the arrow and + signs!

The compounds are calcium oxide, calcium hydroxide and calcium carbonate [1].

Worked example – a step by step approach to answering questions, guiding you through from start to finish

Sample answer

Ammonia, NH_3, is a simple covalent molecule. Graphite is a macromolecule with a giant covalent structure. Compare the volatility and electrical conductivity of these two molecules and draw a dot-and-cross diagram for ammonia. [7]

Ammonia is volatile and has a low boiling point [1]. Graphite is a giant structure so has a high boiling point [1]. Ammonia does not conduct [1] but graphite does [1].

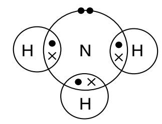

Figure 6.13

Bonding pairs of electrons between each of the three N and H atoms [2].

If this not scored then a pair of bonding electrons between one of the N and H atoms [1].

Lone pair of electrons on the N atom [1].

Sample answer – an example of a question with a good quality answer that is likely to score a high mark.

Exam-style questions

Question 4.01

The electronic structures of the atoms of four elements are:

A 2,8,4 B 2,2 C 2,8 D 2,8,8,1

a Which one of these elements is in Group I of the Periodic Table? Explain your answer. [2]

b Which one of these elements has the lowest proton number? Explain your answer. [2]

c Which one of these elements is in Period 4 of the Periodic Table? Explain your answer. [2]

d Which element is a noble gas? Explain your answer. [2]

e Element A has three naturally-occurring isotopes

i What is meant by the term *isotope*? [1]

ii An isotope of A has a nucleon number of 30. State the number of electrons, protons and neutrons in this isotope. [3]

Question 4.02

Two isotopes of bromine are $^{79}_{35}Br$ and $^{81}_{35}Br$.

a Deduce the number of neutrons in each of these isotopes. [1]

b Bromine has the electronic structure 2,8,18,7.

i Explain how this structure shows that bromine is in Group VII of the Periodic Table. [1]

ii Explain how this structure shows that bromine has a proton number of 35. [2]

c Magnesium reacts with bromine to form magnesium bromide.

i Write the electronic structure for a bromide ion. [1]

ii Write the electronic structure for a magnesium ion, Mg^{2+}. [1]

Unit 4 Atomic structure

Revision checklist

You should be able to:

- ☐ State the relative charges and approximate relative masses of protons, neutrons and electrons
- ☐ Define nucleon number (mass number) as the total number of protons and neutrons in the nucleus of an atom
- ☐ Explain the basis of the Periodic Table in terms of the number of protons
- ☐ Explain that isotopes are atoms of the same element with different numbers of neutrons
- ☐ Understand that isotopes have the same properties because they have the same number of electrons in their outer shell
- ☐ State the two types of isotopes as being radioactive and non-radioactive
- ☐ State one medical and one industrial use of radioactive isotopes
- ☐ Describe the electronic structure of atoms
- ☐ Understand the importance of the noble gas electronic structure

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Exam-style questions – practice answering exam-style questions and check your answers against those provided at the back of the book

Revision checklist – at the end of each chapter so you can check off the topics as you revise them

Introduction

This book is designed to support students studying the Cambridge IGCSE Chemistry syllabus (0620) from Cambridge International Examinations. The topics in the syllabus have been divided into 30 units and match the syllabus.

The main purpose of this publication is to serve as a revision guide for students. The features of this book outlined above are designed to make learning as effective as possible and to give plenty of opportunity to test yourself and gain confidence before taking examinations. Material indicated by a red line relates to the material that forms the supplementary content of the syllabus.

Practical aspects of chemistry are considered not only in Units 2 and 3 but also throughout the book where relevant. These include questions relevant to practical examinations and to alternatives to practical examinations as well as to coursework.

Particles in motion

Learning outcomes

By the end of this unit you should:

- State the distinguishing properties of solids, liquids and gases
- Describe the structure of solids, liquids and gases in terms of particle separation, arrangement and types of motion
- Be able to describe changes in state
- Explain changes of state in terms of the kinetic theory
- Understand the effect of pressure and temperature on a gas in terms of the kinetic theory
- Describe Brownian motion as evidence for the kinetic theory
- Describe and explain diffusion
- State evidence for Brownian motion
- Know how the rate of diffusion depends on the relative molecular mass of the molecules

1.01 States of matter

The three states of matter are solids, liquids and gases.

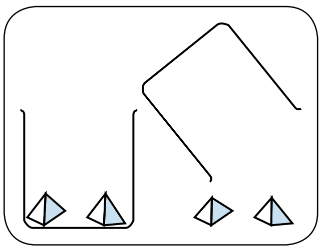
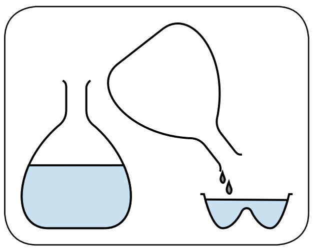
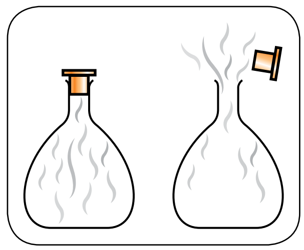
Solid	Liquid	Gas
 Figure 1.01	 Figure 1.02	 Figure 1.03
fixed volume	fixed volume	no definite volume
fixed shape; they keep this shape unless hit	no definite shape	no definite shape
does not flow easily (unless the solid is a powder)	flows easily	spreads out everywhere
cannot be compressed	only compressed a little if at all	can be compressed

Table 1.01

We can explain these bulk properties using the particle theory.

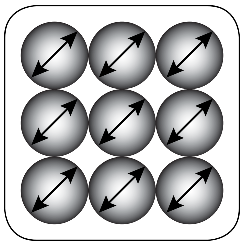
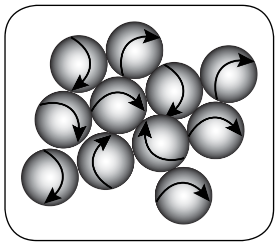
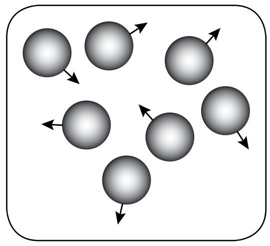
Solid	Liquid	Gas
 Figure 1.04	 Figure 1.05	 Figure 1.06
particles close together	particles close together	particles far apart
particles arranged in a regular pattern	particles arranged randomly	particles arranged randomly
particles vibrate around a fixed point	particles slide over each other randomly and slowly	particles move randomly and rapidly

Table 1.02

TIP

It is a common error to think that the particles in liquids are spaced out. They are not. They are close to each other with very little or no space between. They slide over each other and do not have free motion like gases.

Sample answer

Use the kinetic particle theory to explain why a crystal of iodine keeps its particular shape and cannot be compressed but iodine vapour can be compressed and spreads everywhere. [4]

The particles in solid iodine are regularly arranged [1] so it keeps its shape. They are packed closely together [1] so the crystal cannot be compressed. The particles in iodine vapour are far apart [1] (as there are no attractive forces between them) [1]. When pressure is put on the vapour, the particles can be pushed closer to each other [1].

Progress check

1.01 Describe the three states of matter in terms of shape and volume. [6]

1.02 Describe the difference between solids and liquids in terms of closeness and motion of particles. [4]

1.03 At room temperature and pressure bromine molecules are close together and randomly arranged. Describe the proximity (closeness) and arrangement of bromine molecules in (a) bromine vapour [2] (b) solid bromine. [2]

1.02 Changes of state

$$\text{SOLID} \underset{\text{freezing / solidifying}}{\overset{\text{melting / fusion}}{\rightleftharpoons}} \text{LIQUID} \underset{\text{condensing / liquefying}}{\overset{\text{boiling / evaporation}}{\rightleftharpoons}} \text{GAS}$$

TERMS

Condensing: The change in state from a gas to a liquid.

Evaporation: The change in state from a liquid to a gas below the boiling point of the liquid.

Under given conditions of temperature and pressure some substances, for example carbon dioxide, sublime. They turn directly from solid to gas or gas to solid.

$$\text{SOLID} \underset{\text{sublimation}}{\overset{\text{sublimation}}{\rightleftarrows}} \text{GAS}$$

TERMS

Sublimation: The change in state directly from solid to gas and/or gas to solid without the liquid state being formed.

1.03 The kinetic particle theory

TERMS

Kinetic theory: Particles in solids, liquids and gases behave as hard spheres which are constantly moving from place to place (in liquids and gases) or vibrating (in solids).

- Particles in liquids and gases are in constant motion.
- Particles in a gas are constantly colliding and changing directions. They move randomly.
- As the temperature increases, the particles gain more energy and they move faster.
- The simple **kinetic theory** also assumes that the particles are 'hard' spheres.

We can use the kinetic particle theory to explain many facts, for example:

- Gases can be compressed easily: the particles are far apart. Increasing the pressure on a gas decreases the distance between the particles. So an increase in pressure at constant temperature decreases the volume.

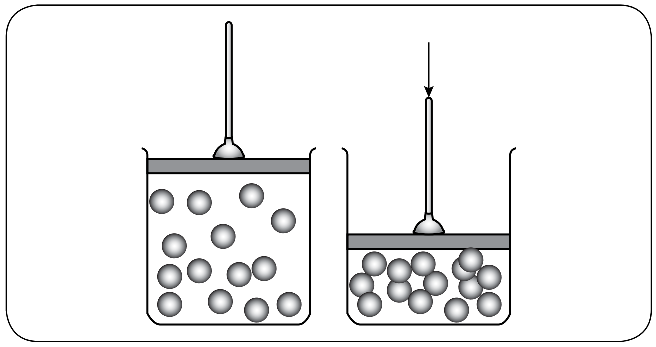


Figure 1.07 Increasing pressure on a gas

- Liquids and solids cannot be compressed easily: the particles are close together or touching. Increasing the pressure has little effect since the particle theory assumes that the particles are 'hard'.
- A gas in a closed container increases in pressure when it is heated. The particles have more energy. They move faster and hit the walls of the container with greater force.
- The volume of gas in a syringe increases when heated at constant pressure: the particles have more energy the higher the temperature, so they hit the walls of the syringe more often pushing the syringe plunger out.

1.04 Explaining changes of state

- An input of energy is needed to melt and boil a substance.
- The energy lessens the forces of attraction between particles when a solid melts or a liquid boils.
- Energy is given out when a substance condenses or freezes.
- This is because the particles are losing some of the energy of their movement (kinetic energy) when condensing or freezing.

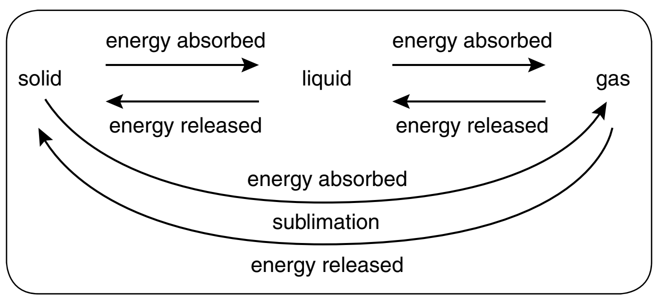


Figure 1.08

TIP

When writing about forces make sure that you:

- refer to attractive forces
- refer to forces between particles not within particles

Worked example 1.01

Use the kinetic particle theory to explain why a balloon increases in size when the temperature increases. [5]

In this type of question your answer should involve a comparison, for example the more energy the higher the temperature (not high temperature). The sequence needed is:
movement → collisions → energy → force → pressure

The particles move faster the higher the temperature [1] and collide more frequently with the walls of the balloon [1]. The gas particles have more energy the higher the temperature is [1] and hit the walls of the balloon with more force [1].

This leads to greater pressure on the wall of the balloon [1].

1.05 Heating and cooling curves

Heating and cooling curves are graphs showing how the temperature changes when a substance is heated or cooled at a steady rate (steady increase or decrease in energy).

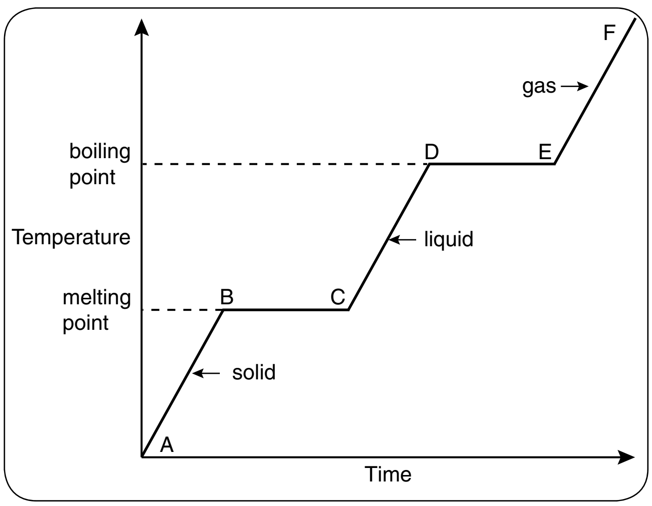


Figure 1.09 A heating curve

A → B: an input of energy increases the temperature of the solid.

B → C: the energy is being used to reduce the attractive forces between the particles in the solid. Liquid as well as solid material is present. The energy does not go in to raise the temperature.

C → D: the energy increases the temperature of the liquid.

D → E: the energy is being used to overcome the attractive forces between the particles in the liquid so that a liquid as well as a gas (vapour) is present.

E → F: the energy increases the temperature of the gas.

The shape of a cooling curve is the reverse of that of a heating curve.

Worked example 1.02

Explain, using ideas of particles and energy, the changes in arrangement and motion of the particles which occurs when zinc melts. [6]

You should make sure that the word particles or a type of particle, for example molecule, is mentioned. Start with the solid arrangement and energy:

The particles in solid zinc are regularly arranged [1] and only vibrate [1]. As temperature increases the zinc particles gain energy [1] and the forces between the zinc particles weaken [1].

Then write about the liquid:

In the liquid particles move over each other [1] and become irregularly arranged [1].

Progress check

1.04 Explain why liquids cannot be compressed. [2]

1.05 Name these changes of state: (a) gas to liquid [1], (b) liquid to gas below the boiling point [1], (c) solid to gas (without liquid being formed) [1].

1.06 Copy and complete: When liquid sulfur is heated, the temperature of the liquid ________. At the ________ point the ________ is used to overcome the ________ forces between the ________ and the temperature remains ________. (Words to use: attractive; boiling; constant; energy; increases; molecules.) [6]

1.06 Brownian motion

Brownian motion provides evidence for the kinetic particle model.

Examples: 1. the zigzag movement of pollen grains in still water

2. random movement of dust particles in still air

TERMS

Brownian motion: The random movement of small visible particles in a suspension caused by the unequal random bombardment of molecules of liquid or gas on the visible particles.

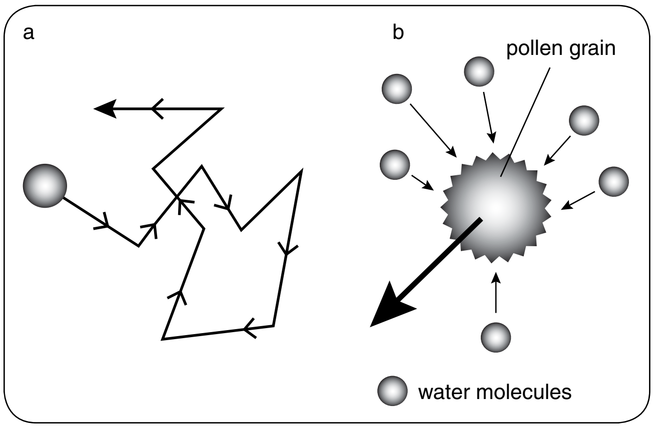


Figure 1.10 **a** Random motion of pollen grain **b** Random bombardment of water molecules on a pollen grain causes movement in the direction →

1.07 Diffusion

TERMS

Diffusion: The spreading movement of one substance into another due to the random motion of the particles.

- Diffusion is explained by the kinetic particle theory. Particles collide randomly with each other leading to complete mixing of the particles.
- Diffusion in gases is faster than diffusion in liquids because the particles in gases move faster.
- Diffusion does not happen in solids because the particles are fixed in position.
- The overall movement of particles is from where they are more concentrated to where they are less concentrated.

TIP

When writing about diffusion refer to the random movement of the particles. It is a mistake to think that the particles always travel in the direction from high to low concentration.

Diffusion of gases

Figure 1.11 shows the diffusion of bromine vapour in air.

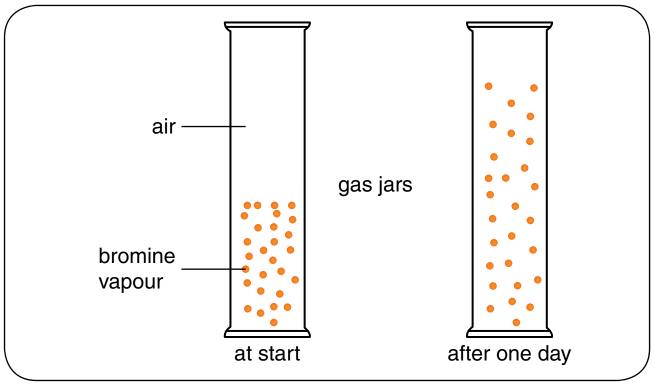


Figure 1.11

Diffusion of liquids

Figure 1.12 shows the diffusion of ink in water.

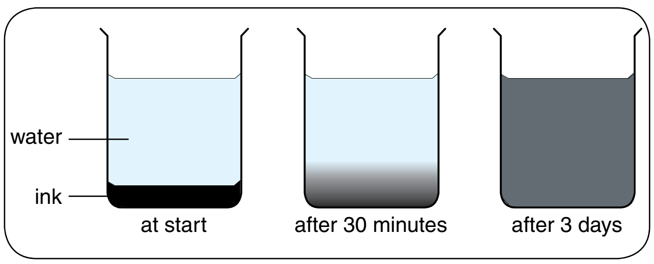


Figure 1.12

Diffusion and relative molecular masses

- The masses of molecules are compared by using relative molecular masses, M_r (see Unit 8).
- The higher the value of M_r the heavier the molecule.
- Molecules with lower M_r move faster than molecules with higher M_r.

Progress check

1.07 Describe the evidence for the kinetic particle theory. [3]

1.08 Define and explain the term diffusion using the kinetic particle theory. [4]

1.09 Explain why hydrogen gas, H_2, diffuses quicker in air than hydrogen sulfide gas, H_2S. [2]

Sample answer

The wall of a rubber balloon is slightly porous. Gases can move through the wall. A balloon contains a mixture of carbon dioxide and helium. Explain why the percentage of carbon dioxide in the balloon increases with time. [4]

Molecules with lower relative molecular mass move faster [1] by diffusion [1] than molecules with higher molecular mass. Helium has a lower molecular mass than carbon dioxide [1]. So helium diffuses faster out of the balloon than carbon dioxide [1].

Exam-style questions

Question 1.01

The arrows on the diagram below represent some changes in state.

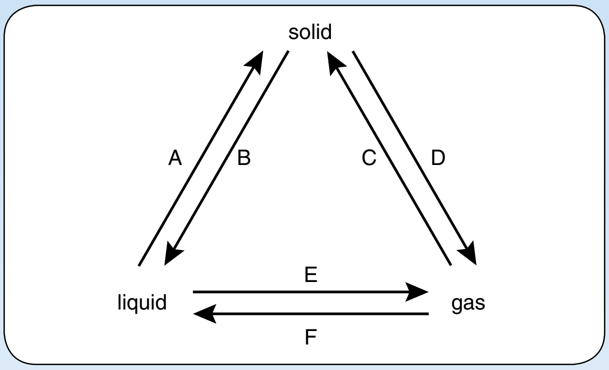


Figure 1.13

a Give the name of the changes B, D and E. [2]

b Describe what is happening to the arrangement and motion of the particles in change A. [4]

c What energy change is taking place in
  i change B [1]
  ii change F? [1]

Question 1.02

Hydrogen chloride, HCl, and hydrogen bromide, HBr, are gases which turn blue litmus paper red. A teacher soaked some cotton wool in hydrochloric acid and set up the apparatus shown below. Hydrogen chloride gas evaporated from the hydrochloric acid.

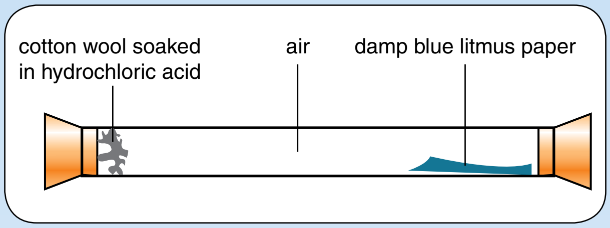


Figure 1.14

Hydrogen chloride gas evaporated from the cotton wool. It was only after two minutes that the litmus paper turned red.

a Use the kinetic particle theory to explain these results. [3]

b The teacher repeated the experiment with hydrobromic acid, which evaporates to produce hydrogen bromide gas. Would the litmus paper turn red, quicker or slower or take the same time? Explain your answer. [3]

c Explain evaporation using the kinetic particle theory. [3]

Question 1.03

Explain using the kinetic particle theory why particles of smoke in still air appear to move in an irregular way. [5]

Revision checklist

You should be able to:

- ☐ State the distinguishing properties of solids, liquids and gases
- ☐ Describe the structure of solids, liquids and gases in terms of particle separation, arrangement and types of motion
- ☐ Describe changes of state in terms of melting, boiling, evaporation, freezing, condensation and sublimation
- ☐ Explain changes of state in terms of the kinetic theory
- ☐ Describe the temperature and pressure of a gas in terms of the kinetic theory
- ☐ Describe Brownian motion as evidence for the kinetic theory
- ☐ Explain Brownian motion in terms of particle collisions
- ☐ Describe and explain diffusion
- ☐ Explain how diffusion depends on relative molecular mass

Experimental chemistry

Learning outcomes

By the end of this unit you should:

- Be able to name and know the use of glassware, for example pipettes, burettes
- Be able to name apparatus for the measurement of time, temperature, mass and volume
- Know that density is the mass of substance in a given volume
- Be able to describe and comment on experimental arrangements
- Know how to record readings, complete tables of data and plot graphs
- Be able to plan investigations, taking into account the control of variables

2.01 Basic laboratory apparatus

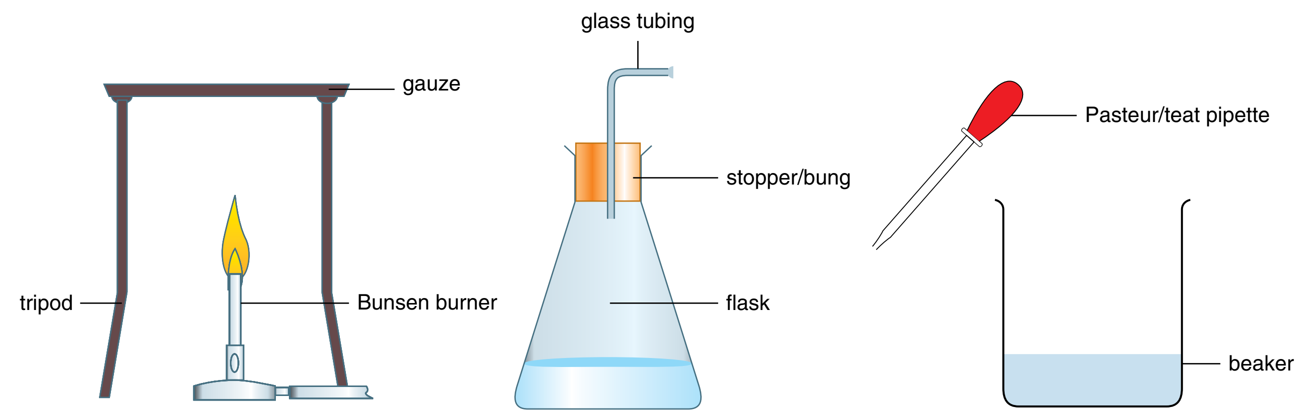


Figure 2.01

TIP

You should draw apparatus in cross section not in three dimensions. Make sure that if tubes come out of flasks, they are not cut across by a line on your diagram.

- Flasks are used for carrying out reactions where fairly small amounts of liquid are used.
- Beakers are used to store liquids temporarily or sometimes for carrying out reactions.
- Test tubes are used to carry out reactions where small amounts of liquid are used and not heated.
- Boiling tubes are used to heat small amounts of liquid.

2.02 Recording mass, time and temperature

Mass

- Mass is recorded on a balance. A good balance can give a reading to two decimal places, for example 45.15.
- The unit of mass is the gram (g) or kilogram (kg). 1000 g = 1 kg
- We change g into kg by dividing the mass in g by 1000.

Time

- Time is recorded on a stop clock or stop watch.
- The unit of time is the second (s). For chemical reactions which are slower, we can use minutes (min).

Temperature

- Temperature is measured with a thermometer.
- The unit of temperature is degree Celsius (°C). We can read accurate thermometers to ±0.1 °C.

2.03 Measuring volumes of liquids

- Volume is measured in centimetres cubed (cm^3) or decimetres cubed (dm^3): $1000\,cm^3 = 1\,dm^3$.
- We can change cm^3 into dm^3 by dividing the volume in cm^3 by 1000. For example, $25\,cm^3 = 25/1000 = 0.025\,dm^3$.
- A burette is used to accurately deliver up to $50\,cm^3$ of liquid. The scale divisions of a burette can be read to the nearest $0.1\,cm^3$.
- A volumetric pipette can deliver a single fixed volume of liquid very accurately, for example $25.0\,cm^3$ of sodium hydroxide solution. Some volumetric pipettes have scale divisions like a burette which can be read accurately.

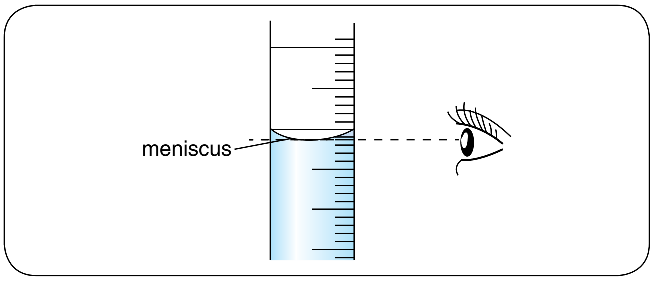


Figure 2.03 When reading a burette, your eye should be level with the bottom of the meniscus

- A volumetric flask is used to make up a solution of a known concentration very accurately. A known mass of solute is dissolved in a small amount of solvent in the flask. The flask is then filled to the graduation mark with solvent.
- A measuring cylinder is used for measuring volumes of solutions where accuracy is not so important. Many measuring cylinders have scale divisions which are only every $2\,cm^3$.

Progress check

2.01 Convert 0.84 kilograms to grams. [1]

2.02 Convert 23 grams to kilograms. [1]

2.03 Convert $36\,cm^3$ to dm^3. [1]

2.04 Convert $0.450\,dm^3$ to cm^3. [1]

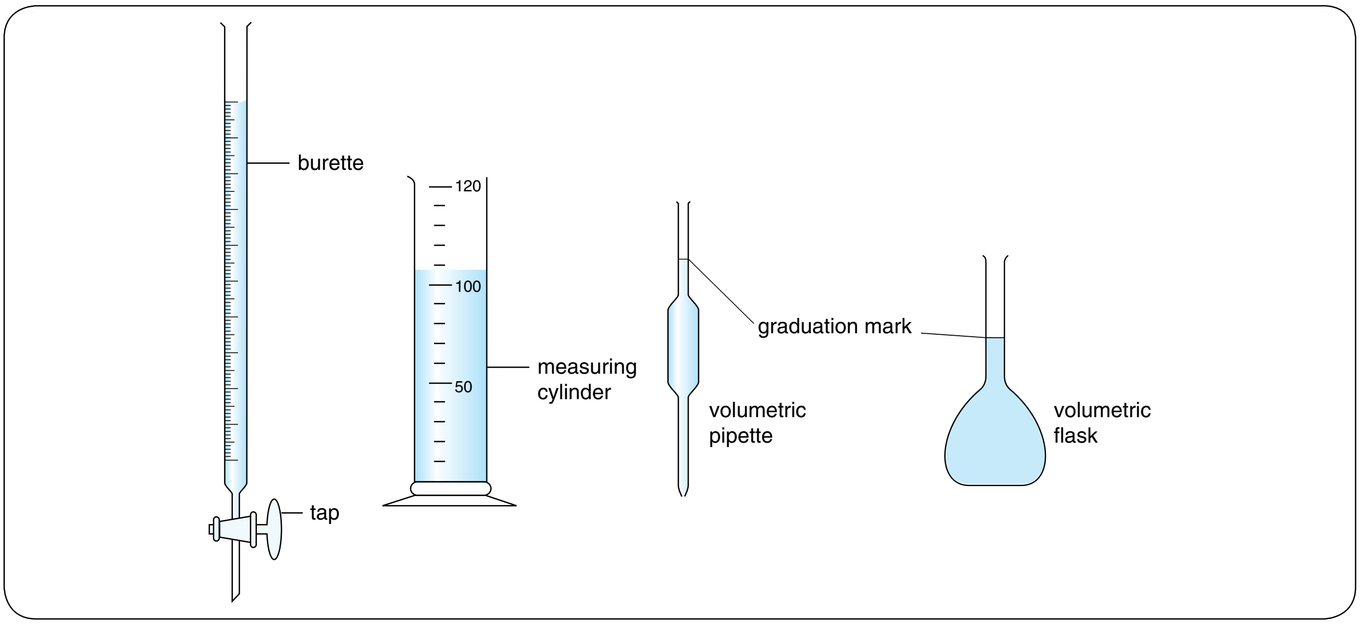


Figure 2.02 Apparatus for measuring volumes of liquids

Sample answer

A student used the apparatus in Figure 2.04 to calculate the concentration of aqueous sodium hydroxide.

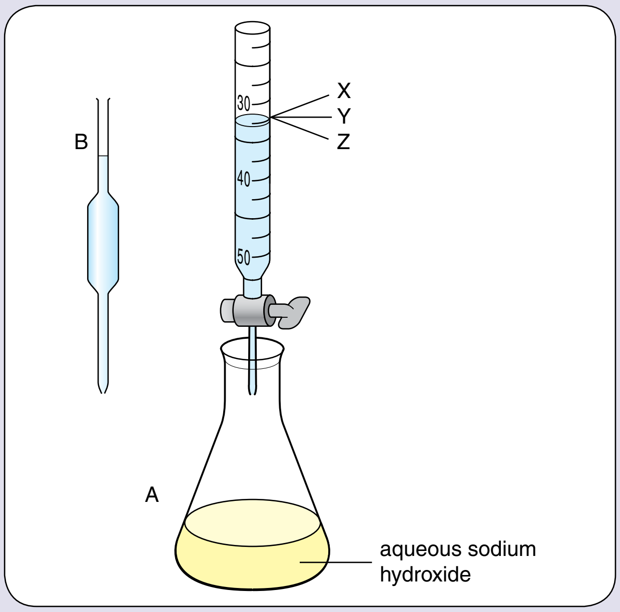


Figure 2.04

a Give the names of the pieces of glassware labelled A and B. **[2]**

A (conical) flask / Erlenmeyer (flask) **[1]**, B volumetric pipette **[1]**.

b In which place, X, Y or Z, should you position your eye to read the burette? Give a reason for your answer. **[2]**

Y **[1]**, your eye should be in line with the bottom of the meniscus **[1]**.

c The diagram shows the initial burette reading and the final burette reading.

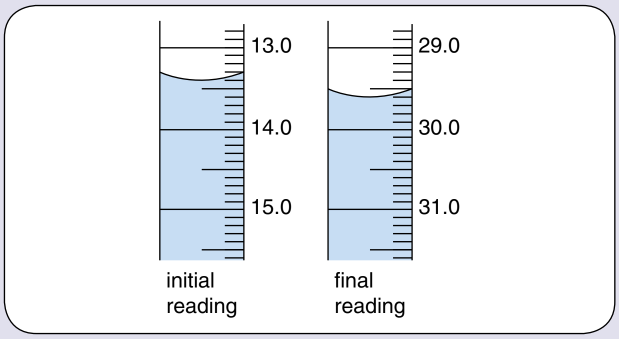


Figure 2.05

Give the initial and final burette readings and deduce the volume of hydrochloric acid which was delivered into the flask. **[3]**

Initial reading = 13.4 (cm^3) **[1]**, final reading = 29.6 (cm^3) **[1]**.

volume = final − initial = 29.6 − 13.4 = 16.2 (cm^3) **[1]**.

Worked example 2.01

When a solution of acid is added to a solution of alkali, the temperature of the mixture increases. A student wants to measure how the temperature varies with the volume of acid added to the alkali.

What equipment should the student use to carry out the experiment accurately and what measurements should be taken? **[8]**

1 *Since temperatures are being measured, heat losses should be minimised:*

Insulated beaker or drinking cup for the alkali **[1]**

2 *Select apparatus for measuring accurate volumes:*

Burette for measuring volumes of acid **[1]**

Fixed volume of alkali put into a beaker measured with a volumetric pipette or burette **[1]**

3 *Carry out the procedure, taking relevant measurements:*

Add known small volume of acid to alkali **[1]**

Stir **[1]**

Record the highest temperature of the solution **[1]**

With an accurate thermometer **[1]**

Repeat by adding further fixed volumes of acid and measuring the temperature **[1]**

2.04 Measuring gas volumes

The volume of a gas produced during a reaction (see Unit 12) can be measured using:

- a gas syringe – **a** in Figure 2.06
- an upturned measuring cylinder (or upturned burette) full of water at the start of the experiment – **b** in Figure 2.06

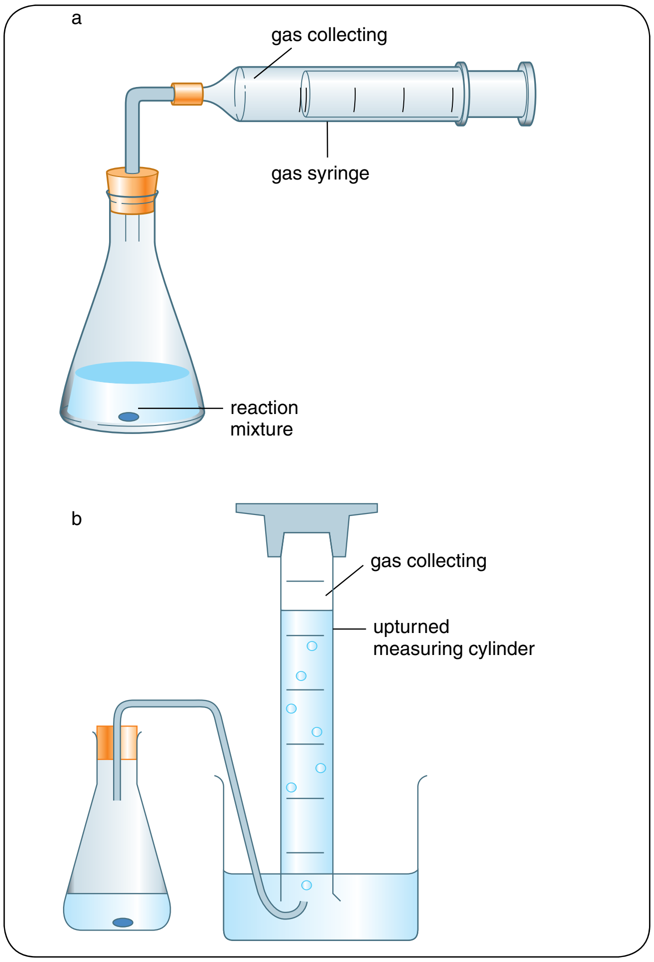


Figure 2.06

Progress check

2.05 What piece of apparatus should you use to measure out 23.4 cm^3 of solution accurately? [1]

2.06 What two different apparatuses you could use for measuring the volume of gases. [2]

2.07 Why would you not use a measuring cylinder to measure 25.0 cm^3 of a solution into a beaker? [2]

2.05 Density

- Density (g/cm^3) $= \dfrac{\text{mass (g)}}{\text{volume (cm}^3\text{)}}$
- Gases with densities that are less than air can be collected by the downward displacement of air – Figure 2.07 **a**.
- Gases with densities that are greater than air can be collected by the upward displacement of air – Figure 2.07 **b**.

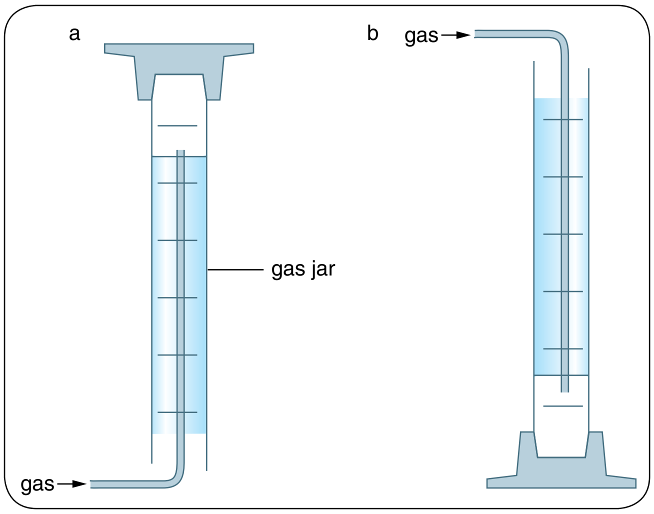


Figure 2.07

- If two liquids do not mix, the less dense one floats on the one that is denser.

TIP

It is better not to use the words 'lightweight' or 'light' when referring to density. They do not indicate mass for **a given volume**.

2.06 Carrying out experiments

When carrying out experiments, we need to decide:

- the apparatus to be used
- the conditions needed, e.g. heat, catalyst
- the measurements to be made
- what variables are involved
- how to make the experiment a fair test
- the accuracy and reliability of the measurements
- safety, for example if harmful gases are given off, use a fume cupboard

Variables and fair testing

A variable is something that is changed in an experiment.

Experiments involving measurements may have several variables.

When calcium carbonate reacts with hydrochloric acid, the volume of gas given off can be measured at various time intervals using the apparatus in Figure 2.06 a.

In this experiment the variables are:

- volume of gas given off
- time
- temperature
- mass of calcium carbonate
- size of calcium carbonate lumps
- concentration and volume of hydrochloric acid

So if we want to find out how volume of gas changes with time:

- We need to keep temperature, the mass of calcium carbonate, the size of the calcium carbonate lumps and the concentration and volume of hydrochloric acid constant (control variables).
- We need to measure the volume of carbon dioxide at different times.
- The variable selected by you (in this case time) is called the independent variable.
- The variable which is measured for each change of the independent variable (in this case the volume of gas) is called the dependent variable.
- The experiment above is a **fair test**.

TERMS

Fair test: An experiment where the independent variable affects the dependent variable and all other variables are controlled.

Accuracy

- Repeat your measurements and take an average, ignoring any **anomalous results**.
- Use instruments with as great an accuracy as possible, e.g. a burette for measuring volumes of liquids accurately instead of a measuring cylinder.
- Read the instruments carefully.
- In an experiment, the overall accuracy depends on the least accurate instrument you are using.

TERMS

Anomalous result: A result or piece of data which does not fit the pattern of the rest of the data.

Progress check

2.08 What is meant by a fair test? [2]

2.09 The table shows how the volume of gas released in a reaction varies with time.

time / s	0	10	20	30	40	50
volume / cm^3	0	5	10	12	20	25

Table 2.01

Which is the anomalous reading? Give a reason for your answer. [3]

Tables and graphs

Tables

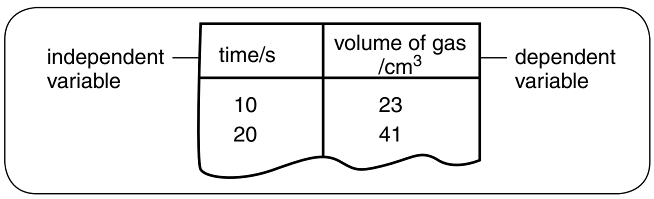


Figure 2.08

Graphs

When drawing graphs:

- plot each point as an ×
- draw lines or curves of best fit
- ignore anomalous points
- the dependent variable is plotted on the y (vertical) axis of the graph

TIP

When the points plotted show a curve, do not join one point to the next by a straight line.

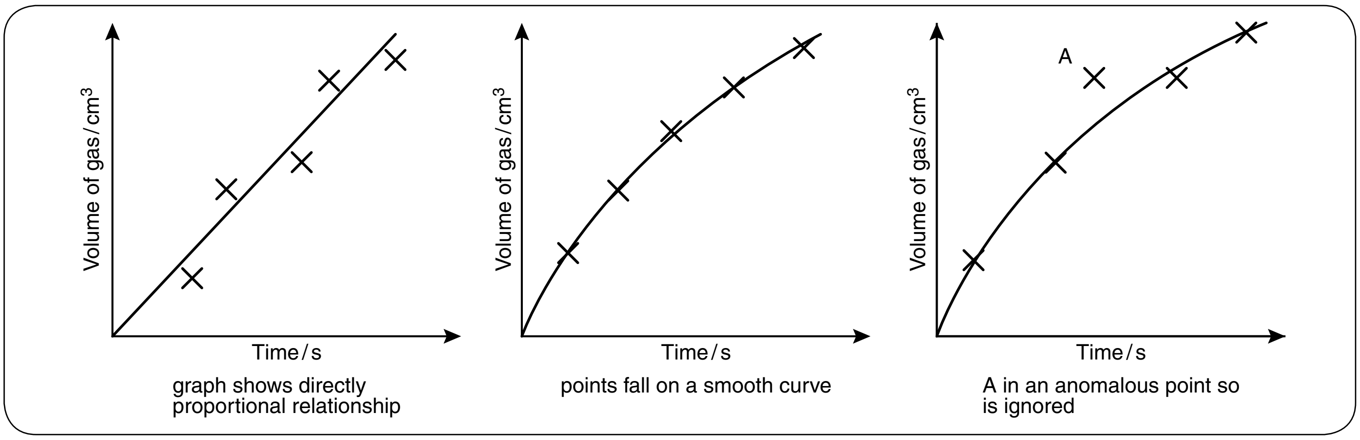


Figure 2.09

Worked example 2.02

When magnesium reacts with hydrochloric acid, hydrogen gas is produced. A student wants to investigate how the volume of gas produced in 20 seconds changes when the temperature changes. Draw the headings of the table that the student should make for the results and list the factors which should be kept constant in this experiment. **[8]**

How to get the answer:

Table:

The quantity chosen by you (independent variable) goes on the left so

Temperature on the left **[1]** with units °C **[1]**

The values you are measuring (dependent variable) goes on the right so

Volume on right **[1]** with units cm^3 **[1]**

Factors kept constant:

The possible variables in an experiment are masses, volumes, concentrations, time, temperature, pressure. You need to select control variables suitable for your experiment. Make sure that you name particular chemicals:

Mass of magnesium **[1]**, size of pieces of magnesium **[1]**

Concentration of hydrochloric acid **[1]**, volume of hydrochloric acid **[1]**

Exam-style questions

Question 2.01

Chlorine is a poisonous gas which is denser than air and soluble in water. A student made chlorine by reacting sodium chloride with concentrated sulfuric acid and collecting the gas using the apparatus in Figure 2.10.

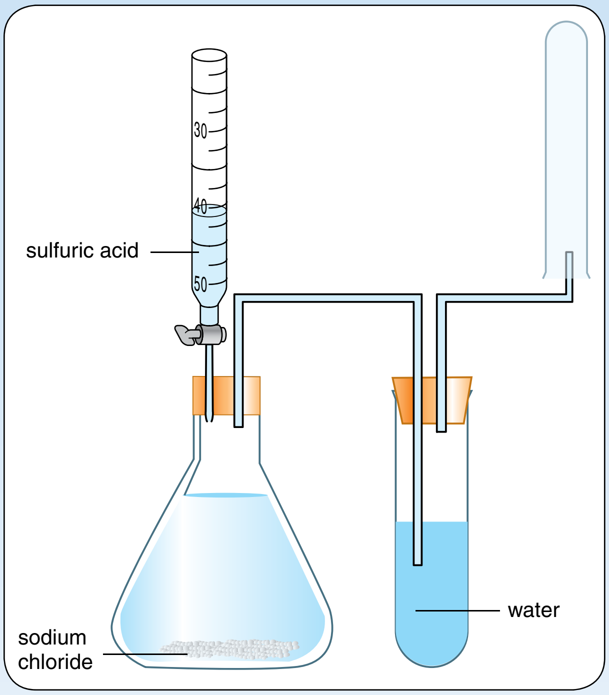


Figure 2.10

a Identify two errors in the diagram. Explain why each is an error. [4]

b Where on the diagram is heat applied? [1]

c Suggest why the experiment should be carried out in a fume cupboard. [1]

d State one hazard of concentrated sulfuric acid. [1]

e State one other safety precaution when carrying out this experiment. [1]

Question 2.02

A student compared the energy released when different fuels were burned in the same spirit burner. The student compared the fuels by measuring the increase in temperature of 100 cm^3 of water in a copper can.

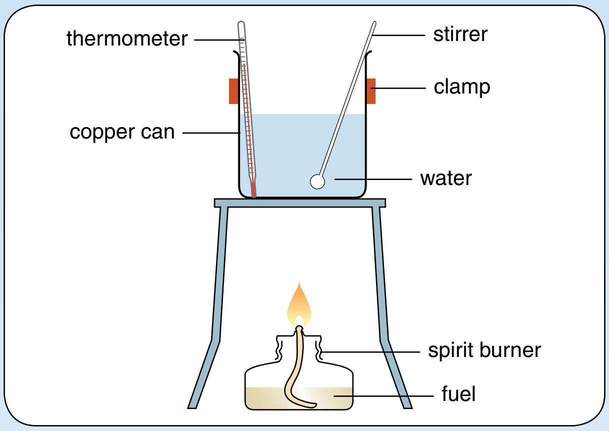


Figure 2.11

a Explain why the student kept the water stirred all the time. [1]

b State two other factors which should be kept constant when comparing the fuels. In each case explain why. [4]

c The student repeated the experiment twice with each fuel. Suggest why. [2]

d What piece of equipment should be used to measure out 100 cm^3 of water? Give a reason for your answer. [2]

e The mass of fuel in the spirit burner was weighed to three decimal places. The reading on the thermometer was to the nearest °C. Why was it of little value to give the answer to the calculation to three decimal places? [1]

Revision checklist

You should be able to:

- ☐ Name appropriate apparatus for the measurement of time, temperature, mass and volume
- ☐ Understand the use and accuracy of burettes, pipettes, measuring cylinders and other laboratory glassware
- ☐ Select suitable apparatus for the accuracy required by the experiment
- ☐ Record readings, construct tables of data and plot graphs
- ☐ Understand the importance of controlling particular variables in an experiment
- ☐ Plan an investigation given relevant information

Unit 3

Methods of purification

Learning outcomes

By the end of this unit you should:

- Understand the importance of purity in substances in everyday life
- Assess the purity of substances from melting and boiling point data
- Describe paper chromatography and interpret simple chromatograms
- Interpret simple chromatograms, including the use of R_f values
- Understand how locating agents are used to make colourless substances visible on a chromatogram
- Describe and explain how solvents, filtration, crystallisation and distillation are used to separate or purify substances
- Suggest suitable purification techniques to obtain a given product

3.01 Pure and impure substances

It is important that substances put into medicines and added to food are pure because the impurities may contain substances which are harmful to us.

- A pure substance is a single substance, for example pure sodium chloride is 100% sodium chloride.
- Impure substances contain other substances mixed with them.
- We can tell the difference between a pure and impure substance by differences in melting points and boiling points.

Pure substance	Impure substance
solid has sharp melting point – all the solid melts at same temperature	solid melts over a temperature range and at a lower temperature than the pure solid
liquid has sharp boiling point – all the liquid boils at the same temperature	1. for a solid dissolved in a liquid: liquid boils over a temperature range and at a higher temperature than the pure liquid 2. for a mixture of two liquids: the mixture starts boiling at the boiling point of one of the liquids and rises to the boiling point of the other liquid

Table 3.01

Progress check

3.01 A solid melts between 234–240 °C. Is this solid likely to be pure or impure? Give a reason for your answer. [1]

3.02 Why should zinc oxide used in creams to treat sunburn be pure? [1]

3.03 Explain why salt is spread on roads in icy weather. [3]

3.02 Chromatography

TERMS

Paper Chromatography: The separation of a mixture of soluble compounds using chromatography paper and a solvent.

Paper chromatography is used to separate and purify a mixture of dissolved substances, for example the pigments (colourings) present in food colourings and inks. Figure 3.01 shows how to carry out paper chromatography.

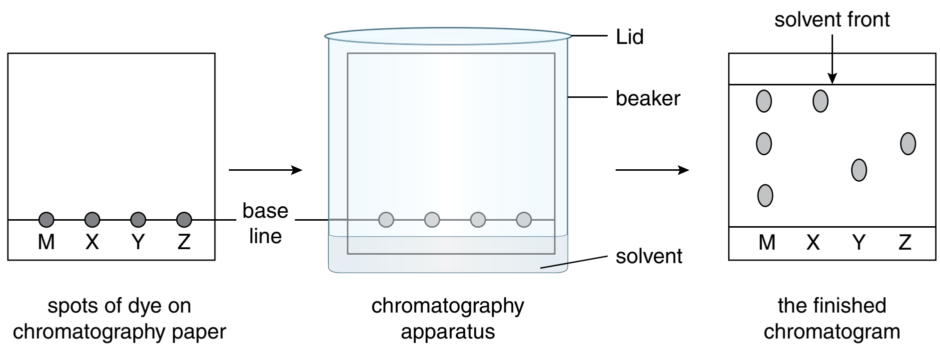


Figure 3.01

- Draw a pencil line on the chromatography paper.
- Put a spot of the concentrated mixture to be separated, M, on the line using a very thin tube.
- Put spots X, Y and Z that you think are in the dye mixture on the base line for comparison.
- Dip the bottom of the paper in the solvent.
- The solvent moves up the paper and allows separation of the mixture.
- Remove the paper when the solvent front is near the top.

The spots on the chromatogram can be compared with those of known dyes. In Figure 3.01, the mixture, M, contains dyes X and Z as well as a third dye.

We can identify the components in the mixture because for a particular solvent they travel a certain distance compared with the solvent front.

TIP

When carrying out chromatography, the solvent level should be below the base line to prevent the mixture washing off into the solvent.

Retention factor (R_f)

TERMS

R_f: In chromatography, the distance moved by a particular substance from the base line divided by the distance moved by the solvent front from the base line.

$$R_f = \frac{\text{distance from base line to centre of spot}}{\text{distance of solvent front from base line}}$$

R_f values can be used to identify compounds or ions because most substances have characteristic R_f values.

Locating agents

TERMS

Locating agent: A substance that reacts with colourless spots on a chromatogram to make them visible as coloured spots.

Many compounds, for example amino acids, are colourless and so cannot be seen on a chromatogram. Spraying the chromatogram with ninhydrin and warming makes the spots appear a purple colour. Ninhydrin is an example of a **locating agent**.

TIP

You do not have to know the names of particular locating agents.

Worked example 3.01

Plant leaves contain pigments called chlorophylls. Describe how you would make a solution of these pigments and use chromatography to separate them. [8]

Step 1 is to extract the pigments:

Grind up leaves **[1]**

With solvent **[1]**

Using a mortar and pestle / using a blender **[1]**

Step 2 is to separate the solids from the pigment solution

Filter off the solid remains of the leaves (through glass wool) **[1]**

Step 3 is chromatography of the pigments

Place a small spot of the solution obtained / filtrate on chromatography / filter paper **[1]**

Place the bottom of the paper in solvent / water / alcohol **[1]**

So that the solvent level is below the spot **[1]**

Allow the solvent to run up the paper (and separate the pigments) **[1]**

Progress check

3.04 In chromatography, why is the base line drawn in pencil and not in ink? [1]

3.05 Draw a diagram of the apparatus used in chromatography. In your diagram include the position of a spot of mixture on the base line and the level of the solvent. [3]

3.06 After chromatography, a spot of dye is 18 cm from the base line. The solvent front is 45 cm from the base line. Deduce the R_f value of the dye. [1]

3.03 Purification by use of a solvent

Liquids which mix with each other are said to be miscible.

Liquids which do not mix are immiscible.

Some substances dissolve in water. Others dissolve better in organic solvents like hexane. We can use these differences in solubility to separate two solutes dissolved in a solvent.

Example:

The solvents hexane and water are immiscible. Iodine is more soluble in hexane than in water. Sodium chloride (salt) is more soluble in water than in hexane. We can separate a mixture of iodine and salt in water by:

1. putting the mixture in a **separating funnel**

TERMS

Separating funnel: A piece of glassware used to separate two immiscible liquids or a solute which is more soluble in one liquid than another.

2. adding hexane to the separating funnel
3. mixing the solutions by shaking
4. allowing the layers to separate on standing. The iodine moves to the hexane layer and the salt remains in the water
5. running off the layer of salt in water. This leaves the iodine in the hexane layer

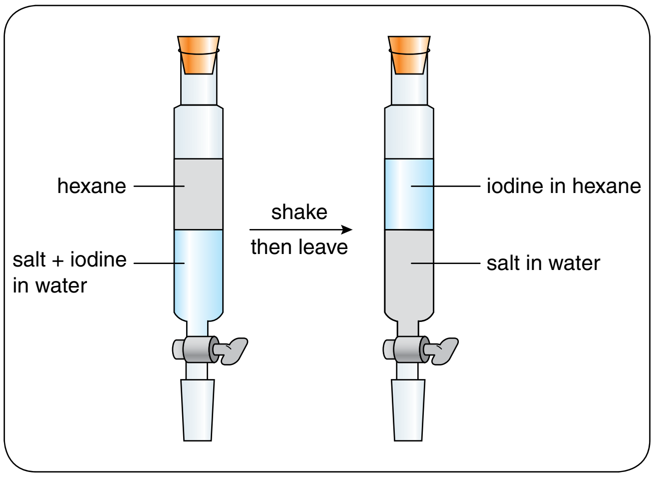


Figure 3.02

Sample answer

Cyclohexane is a volatile solvent. Potassium iodide is soluble in water but not in cyclohexane. Iodine dissolves in both cyclohexane and a solution of potassium iodide in water.

Describe in detail how you would obtain solid iodine from a solution of iodine and potassium iodide in water using cyclohexane. **[8]**

Add cyclohexane to the mixture [1]. Shake the mixture [1] in a separating funnel [1] then allow the layers to settle [1]. Some of the iodine goes into the cyclohexane layer [1] but potassium iodide does not [1]. Then remove the layer of cyclohexane [1] and allow the cyclohexane to evaporate [1] (This leaves the solid iodine).

3.04 Filtration

Filtration separates undissolved solids from a liquid. Molecules of liquid and dissolved substances can flow through the holes in the filter paper but larger particles of solid are too big to pass through. They remain on the filter paper. An example is separating sand from seawater.

- The **filtrate** is the solution passing through the filter paper.
- The **residue** is the solid remaining on the filter paper.
- Traces of solution remaining between particles of solid are removed by washing with a suitable solvent.

TERMS

Filtrate: In filtration, the liquid which goes through the filter paper.

Residue: In filtration, the solid that is trapped on the filter paper.

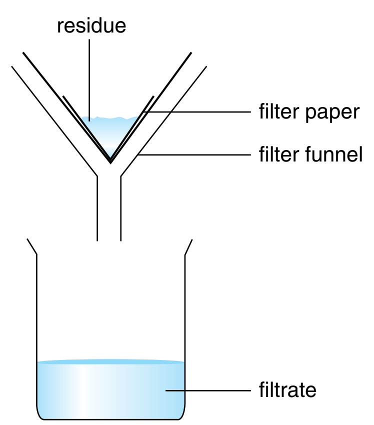


Figure 3.03

3.05 Crystallisation

Crystallisation separates a soluble solid from a solution, for example copper(II) sulfate from its aqueous solution:

1. Gently heat the solution in an evaporating basin to concentrate it.
2. Evaporate the solvent until the crystallisation point (a drop of the solution forms crystals when placed on a cold tile).
3. Leave the saturated solution to form crystals.
4. Filter off the crystals then dry them between filter papers.

3.06 Simple distillation

Simple distillation is used to separate a **volatile** liquid from a solution of a non-volatile solid, for example to separate copper(II) sulfate and water from an aqueous solution of copper(II) sulfate (Figure 3.04).

TERMS

Simple distillation: A method of separating a volatile from a non-volatile substance by evaporation and condensation.

Volatile: Easily changed to a vapour. Volatile substances have low boiling points.

Distillation involves boiling and condensation. It works because the components to be separated have different boiling points.

To separate water from copper(II) sulfate by simple distillation:

1. The solution of copper(II) sulfate in water is heated in a distillation flask.
2. The water boils first because it is volatile. The steam turns to liquid in the condenser.
3. The copper(II) sulfate remains in the distillation flask because it is not volatile – it has a much higher boiling point than water.

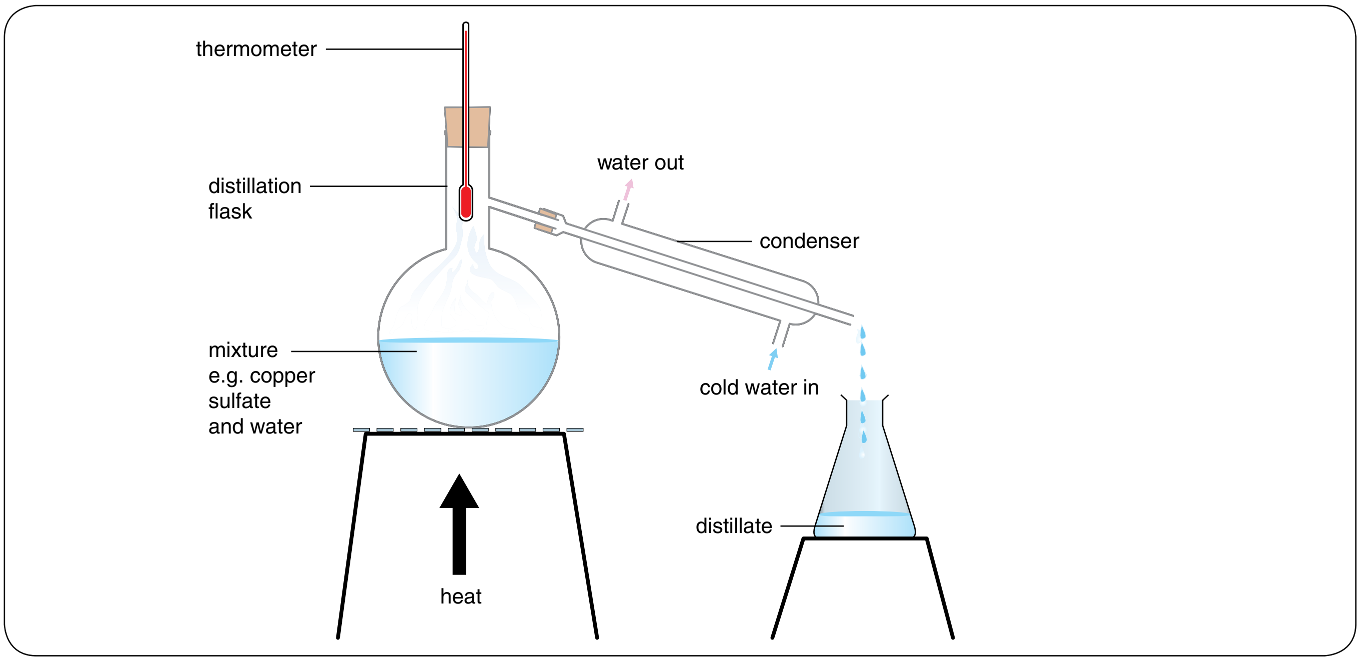


Figure 3.04 Apparatus for simple distillation

3.07 Fractional distillation

TERMS

Fractional distillation: The separation of liquids with different boiling points from a mixture of liquids by evaporation and condensation in a long column.

Fractional distillation is used to separate miscible liquids with different boiling points. It is used to separate petroleum fractions (see Unit 27) and to purify alcohol from a mixture of water and alcohol (Figure 3.05).

- In fractional distillation there is a temperature gradient in the column: higher at the bottom and lower at the top.
- The more volatile components in the mixture (lower boiling points) move faster up the column.
- The less volatile components in the mixture (higher boiling points) do not move as fast.
- The components of the mixture reach the condenser in turn, the most volatile first. They change from vapour to liquid in the condenser and the fractions containing particular components of the mixture are collected one at a time.

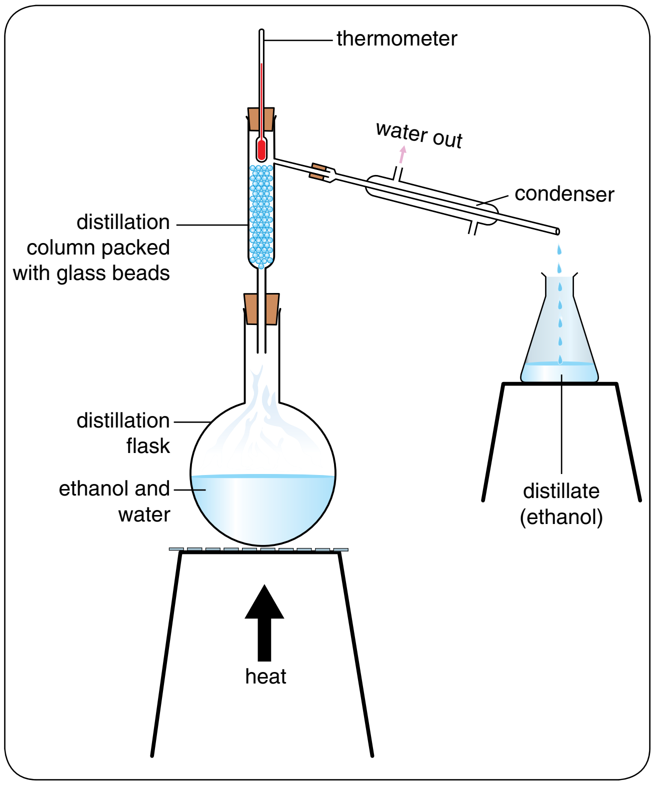


Figure 3.05

3.08 Which purification technique?

To chose the best method to purify a substance from a mixture you have to know the following properties of the different substances in the mixture:

- state at room temperature and pressure
- solubility in water or other solvent
- differences in boiling points

Mixture requiring separation	**Separation method**
insoluble solid and liquid	filtration
crystals of solid from a solution	crystallisation
two soluble solids with different solubility in water and organic solvent	using an organic solvent and separating funnel
a volatile liquid from a solution of a non-volatile solid	simple distillation
a mixture of liquids with different boiling points	fractional distillation

Table 3.02

Worked example 3.02

Describe how you would separate a mixture of sand and aqueous copper(II) sulfate to obtain purified sand, copper(II) sulfate and water. Give an explanation for each stage you use. **[7]**

In order to choose the correct method of separation you need to know the solubility of each component, for example sand is insoluble in water and copper(II) sulfate is soluble in water. So the first step is to separate the solid from the solution.

Filter off the sand **[1]**. Sand is solid but copper(II) sulfate is in solution **[1]**.

You then need to separate the component of the solution. Distillation is appropriate because water has a low boiling point and copper(II) sulfate has a high boiling point.

The filtrate is a solution of copper(II) sulfate **[1]**. This is distilled **[1]**. The copper(II) sulfate remains in the flask **[1]** because it has a high boiling point **[1]**. The water condenses in the condenser and is collected as a distillate **[1]**.

NOTE: simple distillation is used rather than crystallisation, because purified water was required.

Progress check

3.07 What is the best method to separate water from aqueous copper(II) sulfate? [1]

3.08 Hexane and octane are liquids with slightly different boiling points. Suggest the best method of separating these liquids. [1]

3.09 What is the simplest method of separating powdered chalk from water? [1]

3.10 What is the best method of separating two dyes with different solubility in water? [1]

Exam-style questions

Question 3.01

Liquid A has a boiling point of 74 °C. Liquid B has a boiling point of 49 °C.

a Describe with the aid of a diagram how you would obtain purified samples of liquid A and B from a mixture of liquids A and B. [6]

b Liquid B dissolves in both water and pentane. Pentane has a boiling point of 36 °C. Which is the most volatile, B or pentane? Give a reason for your answer. [1]

c The sample of liquid A is impure. What test would you do to show that it is impure and what result would you expect? [2]

d Liquid B reacts with sodium to form a crystalline solid, C. Describe how you could make pure dry crystals of C from an aqueous solution of C. [4]

Question 3.02

Chromatography is used to separate a mixture of carbohydrates which are soluble in an organic solvent. The result is shown in the diagram (Figure 3.06).

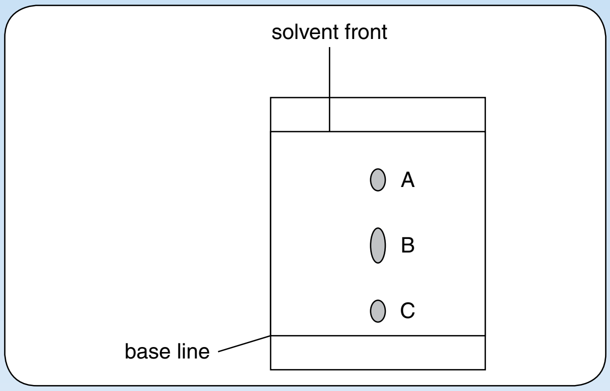


Figure 3.06

a Carbohydrates are colourless compounds. Suggest how the spots can be made visible. [2]

b A student suggested that there were three different carbohydrates in the mixture. What information in the diagram suggests this? [1]

c Why can we not be absolutely certain that there are just three carbohydrates in the mixture? [2]

d Calculate the R_f value of carbohydrate A. [2]

e Which carbohydrate was least soluble in the organic solvent? Give a reason for your answer. [1]

Revision checklist

You should be able to:

- ☐ Describe how paper chromatography is used to separate mixtures
- ☐ Deduce R_f values from chromatograms
- ☐ Understand the purpose of locating agents in making colourless spots visible
- ☐ Understand the importance of purity of substances
- ☐ Identify substances and assess their purity from melting point and boiling point information
- ☐ Describe and explain methods of purification by the use of a suitable solvent, filtration, crystallisation and distillation
- ☐ Suggest suitable purification techniques, given information about the substances involved

Unit 4

Atomic structure

Learning outcomes

By the end of this unit you should:

- State the relative charges and approximate relative masses of protons, neutrons and electrons
- Be able to define proton number and nucleon number
- Use proton number and the simple structure of atoms to explain the basis of the Periodic Table
- Be able to define isotopes
- Know why isotopes of the same element have the same chemical properties
- State the two types of isotopes as being radioactive and non-radioactive
- Know some medical and industrial uses of radioactive isotopes
- Be able to describe the electronic structure of atoms and the significance of the noble gas electronic structures

4.01 Atomic structure

TERMS

Atom: The smallest part of an element that can take part in a chemical change.

- **Atoms** contain sub-atomic particles called protons, neutrons and electrons.
- In the middle of each atom is a tiny nucleus containing protons and neutrons.
- Outside each atom, the electrons are arranged in electron shells (Figure 4.01).

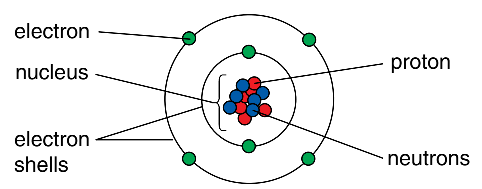


Figure 4.01

The table shows some information about the sub-atomic particles.

Sub-atomic particle	Symbol for the particle	Relative mass	Relative charge
proton	p	1	+1
neutron	n	1	no charge
electron	e^-	0.00054	-1

Table 4.01

4.02 Proton number and the Periodic Table

TERMS

Proton number (atomic number): The number of protons in the nucleus of an atom.

- The number of protons in the nucleus of an atom is called the **proton number (atomic number)**.
- Each atom of the same element has the same number of protons.
- The atoms of the elements are arranged in order of proton number. Hydrogen has one proton, helium has two, lithium three, beryllium four and so on.

- In a neutral atom, the number of electrons = the number of protons.
- The Periodic Table is an arrangement of elements in order of increasing proton number so that elements with the same number of electrons in their outer shell fall in the same vertical column (Group). So Group I elements have 1 electron in their outer shell, Group II elements have 2 electrons in their outer shell and so on.

Sample answer

Describe the structure of a helium atom. Use your Periodic Table to help you.

In your answer include the type, number and position of each sub-atomic particle present. [5]

Helium has a nucleus containing protons and neutrons [1]. There are two protons [1] and two neutrons [1]. The two electrons [1] spin round in a shell outside the nucleus [1].

4.03 Isotopes

TERMS

Isotopes: Atoms of the same element which have the same proton number but a different nucleon number.

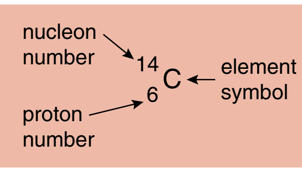
Nucleon number (mass number): The total number of protons and neutrons in the nucleus of an atom.

- Atoms of the same element can have different numbers of neutrons. These atoms are called **isotopes**.
- The total number of protons and neutrons in the nucleus of an atom is called the **nucleon number (mass number)**.
- So isotopes are atoms with the same number of protons but different numbers of neutrons.

TIP

Remember that isotopes refer to atoms not elements. For example, in a molecule with more than one chlorine atom, there may be more than one isotope of chlorine present.

Isotopes can be written with their full name, for example carbon-14, uranium-235. The number after the name is the nucleon number. We usually describe isotopes using standard notation. This shows the chemical symbol, the nucleon number and the proton number.



Three isotopes of hydrogen are shown in Figure 4.02 with their standard notation.

How many neutrons?

Number of neutrons = nucleon number − proton number

So using the notation for chromium with a nucleon number of 52:

$${}^{52}_{24}\mathrm{Cr}$$

There are 24 protons and (52 − 24) = 28 neutrons.

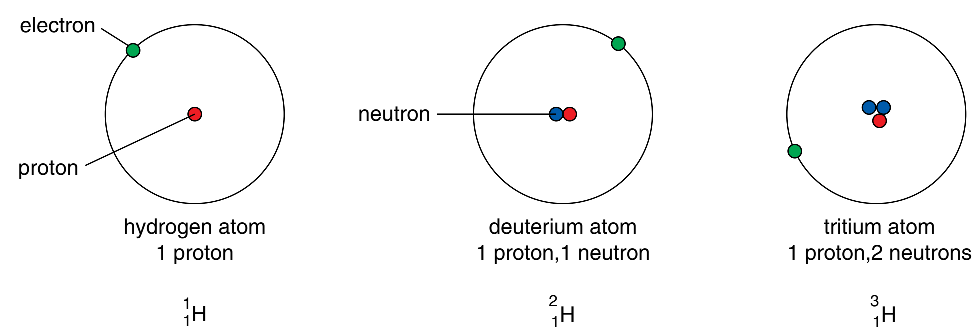


Figure 4.02

TIP

The nucleon number is usually shown on the Periodic Table as well as the proton number. The nucleon number is always the larger of the two figures.

Properties of isotopes

Isotopes of the same element have the same chemical properties because they have the same number of electrons in their outer shell. They may have slightly different physical properties.

Progress check

4.01 Give the relative charges and relative masses of a proton, a neutron and an electron. [6]

4.02 What are *isotopes*? [1]

4.03 Deduce the number of neutrons in an isotope of iron with a nucleon number of 54. [1]

4.04 Deduce the total number of neutrons in a molecule of sulfur $^{32}_{16}S$, which contains 8 sulfur atoms. Show all your working. [2]

4.04 Radioactivity and its uses

Some isotopes are non-radioactive. Others are radioactive. Radioactive isotopes have unstable nuclei. The nuclei break down (decay). As it decays, the nucleus gives out particles or rays. The measurement of the rate of emission of these particles or rays has found many uses.

- Medical uses: cancer treatment, treatment for an overactive thyroid gland, generation of electric current in heart pacemakers, sterilising medical equipment, location of tumours.
- Industrial uses: measuring and controlling the thickness of paper, measuring fluid flow and locating leaks in pipelines, measuring engine wear, energy generation in nuclear power stations.

TIP

The syllabus only requires you to know one medical and one industrial use of radioactive isotopes. Make sure that you know the difference between medical and industrial uses.

Worked example 4.01

Complete the table to show the number of sub-atomic particles in these atoms. [4]

How to get the answer:

Atom	Number of protons	Number of neutrons	Number of electrons
$^{127}_{53}I$	*this is given by the number at the lower left* = 53	*number of neutrons is top left number (nucleon number) minus number of protons (127 – 53)* = 74	*number of electrons in an atom = the number of protons* = 53
$^{176}_{71}Lu$	*this is given by the number at the lower left* = 71	*number of neutrons is top left number (nucleon number) minus number of protons (176 – 71)* = 105	*number of electrons in an atom = the number of protons* = 71

NOTE: 1 mark is given for the first column being correct, 1 mark for the last column being correct and 1 mark each for the number of neutrons.

Table 4.02

Sample answer

Tritium is an isotope of hydrogen. It is radioactive and has a nucleon number of 3.

a What is meant by the term radioactive? [1]

It gives off radiation or particles from unstable atoms [1].

b Describe the similarities and differences in the atomic structure of tritium and ordinary hydrogen. [4]

Both have one proton [1] and one electron [1]. Tritium has 2 neutrons [1] but hydrogen does not have any [1].

c Suggest one industrial use of a radioactive isotope. [1]

Measuring the thickness of paper [1].

4.05 Electron shells

TERMS

Electron shells: The regions at different distances from the nucleus where one or more electrons are found.

In a simple model of the atom the electrons are arranged in orbits around the nucleus. These orbits are called **electron shells** (Figure 4.03).

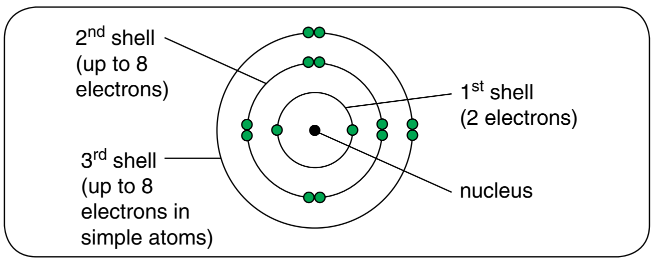


Figure 4.03

- The first shell, nearest the nucleus holds a maximum of two electrons.
- The second shell, further away from the nucleus, holds a maximum of eight electrons.
- The third shell is even further away. It starts filling up when the second shell has eight electrons.
- The fourth shell starts filling up when the third shell has eight electrons.
- Up to calcium (proton number = 20), the shells fill in order 1,2,3,4.

4.06 Deducing electronic structures

TERMS

Electronic arrangement (electronic structure): The number and arrangement of electrons in the electron shells of an atom.

The **electron arrangement** in an atom (also called the electron configuration or electronic structure) is deduced by adding electrons, one at a time, to the electron shells starting with the first shell.

We can write the number of electrons in each shell as a number separated by commas. For example:

- A hydrogen atom has 1 proton, so has 1 electron. This goes into the 1st shell. So the electron arrangement is 1.
- A lithium atom has 3 protons, so has 3 electrons. Two electrons go into the first shell. This shell is then full, so the 3rd electron goes into the 2nd shell. So the electron arrangement is 2,1.
- A sodium atom has 11 protons, so has 11 electrons. Two electrons go into the 1st shell and 8 electrons into the second. The second shell is full, so the 11th electron goes into the 3rd shell. So the electron arrangement is 2,8,1.
- Figure 4.04 shows the electronic structures of some atoms.

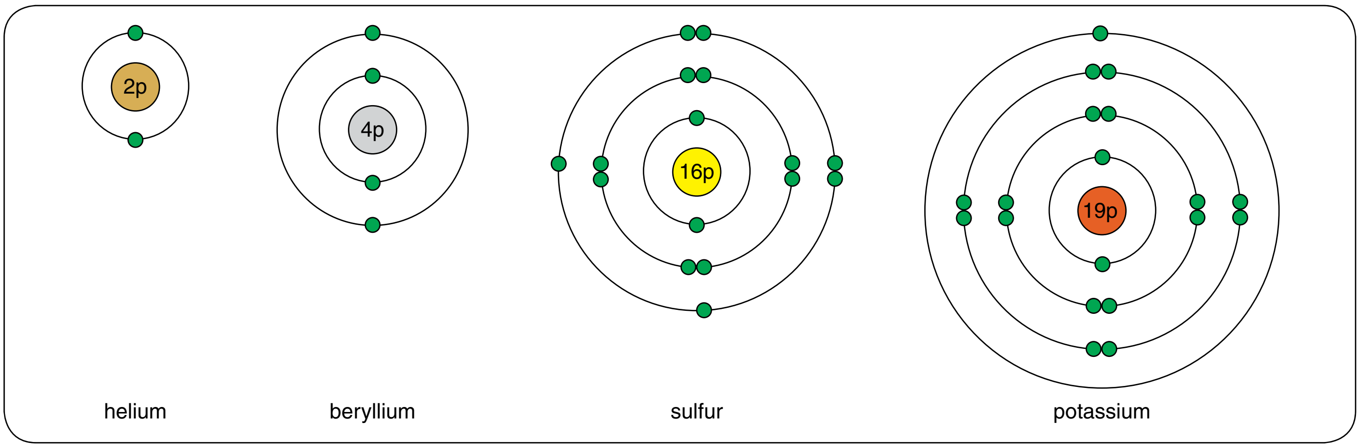


Figure 4.04

The electronic structures of the first 20 elements are:

Element	Number of electrons	Electronic structure	Element	Number of electrons	Electronic structure
hydrogen	1	1	sodium	11	2,8,1
helium	2	2	magnesium	12	2,8,2
lithium	3	2,1	aluminium	13	2,8,3
beryllium	4	2,2	silicon	14	2,8,4
boron	5	2,3	phosphorus	15	2,8,5
carbon	6	2,4	sulfur	16	2,8,6
nitrogen	7	2,5	chlorine	17	2,8,7
oxygen	8	2,6	argon	18	2,8,8
fluorine	9	2,7	potassium	19	2,8,8,1
neon	10	2,8	sodium	20	2,8,8,2

Table 4.03

TIP

You should be able to deduce the electronic structures of the first 20 elements from the information in the Periodic Table because in a neutral atom the number of protons = the number of electrons.

4.07 Electronic structures and the Periodic Table

- The number of outer shell electrons corresponds with the Group number, for example Group III elements have three outer shell electrons and Group VII elements have seven outer shell electrons. The electrons in the outer shell of an atom are sometimes called the valency electrons.
- The number of shells containing electrons gives the Period number.
- The number of electrons in the outer shell determines the chemical properties of an element. Group I metals are very reactive because they can easily lose their outer shell electron.
- The elements in Group I have similar chemical properties because they have the same number of outer shell electrons. The same goes for some other Groups, for example Group II and Group VII.
- The Group VIII elements (noble gases) have a 'full' outer shell of electrons, the maximum number of

Worked example 4.02

a Draw the electronic structures of a chlorine atom and a chloride ion, Cl^-. [3]

Use your Periodic Table to find out the number of electrons in a chlorine atom. This is the same as the proton number (17). Then fill the shells starting from the lowest until you get this number until you get to 17. A chloride ion has one more electron than a chlorine atom.

b Explain in terms of electronic structure why chloride ions and bromide ions are both stable and have similar chemical properties. [5]

There are two parts to this question: the stability and the chemical properties. Make sure that you answer both. The stability relates to the noble gas structure and the reactivity to the number of electrons in the outer shell.

Chlorine and bromine are in Group VII so the electronic structure of their ions is similar [1]. Both ions have eight electrons in their outer shell [1]. This is the noble gas structure [1] which is stable [1]. Since both ions have the same number of electrons in the outer shell they have similar chemical properties [1].

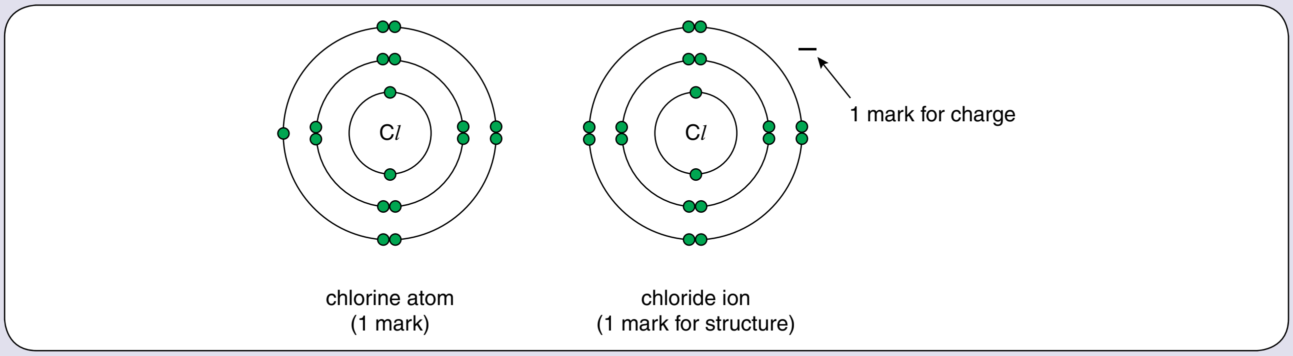


Figure 4.05

electrons the shell can hold (two for helium and eight for the others). This is called the **noble gas structure**.

- A full outer shell of electrons makes these atoms stable because it takes a lot of energy to remove an electron from a full shell.

TERMS

Noble gas structure: The electronic structure of ions or atoms with a complete outer shell of electrons.

Progress check

4.05 Write the electronic structure of

a a silicon atom

b a calcium atom

c a fluorine atom. [3]

4.06 How many electrons are in the second electron shell, when the third shell starts filling? [1]

4.07 a Write the electronic structures of an aluminium ion, Al^{3+} and an oxide ion, O^{2-}. [2]

b Explain why these ions are stable. [2]

Exam-style questions

Question 4.01

The electronic structures of the atoms of four elements are:

A 2,8,4 B 2,2 C 2,8 D 2,8,8,1

a Which one of these elements is in Group I of the Periodic Table? Explain your answer. [2]

b Which one of these elements has the lowest proton number? Explain your answer. [2]

c Which one of these elements is in Period 4 of the Periodic Table? Explain your answer. [2]

d Which element is a noble gas? Explain your answer. [2]

e Element A has three naturally-occurring isotopes

i What is meant by the term *isotope*? [1]

ii An isotope of A has a nucleon number of 30. State the number of electrons, protons and neutrons in this isotope. [3]

Question 4.02

Two isotopes of bromine are $^{79}_{35}Br$ and $^{81}_{35}Br$.

a Deduce the number of neutrons in each of these isotopes. [1]

b Bromine has the electronic structure 2,8,18,7.

i Explain how this structure shows that bromine is in Group VII of the Periodic Table. [1]

ii Explain how this structure shows that bromine has a proton number of 35. [2]

c Magnesium reacts with bromine to form magnesium bromide.

i Write the electronic structure for a bromide ion. [1]

ii Write the electronic structure for a magnesium ion, Mg^{2+}. [1]

Revision checklist

You should be able to:

- ☐ State the relative charges and approximate relative masses of protons, neutrons and electrons
- ☐ Define nucleon number (mass number) as the total number of protons and neutrons in the nucleus of an atom
- ☐ Explain the basis of the Periodic Table in terms of the number of protons
- ☐ Explain that isotopes are atoms of the same element with different numbers of neutrons
- ☐ Understand that isotopes have the same properties because they have the same number of electrons in their outer shell
- ☐ State the two types of isotopes as being radioactive and non-radioactive
- ☐ State one medical and one industrial use of radioactive isotopes
- ☐ Describe the electronic structure of atoms
- ☐ Understand the importance of the noble gas electronic structure

Unit 5

Elements, compounds and mixtures

Learning outcomes

By the end of this unit you should:

- Know the difference between physical and chemical changes
- Be able to describe the differences between elements, compounds and mixtures
- Be able to describe the differences between metals and non-metals
- Be able to describe an alloy as a mixture of a metal with other elements

5.01 Physical and chemical changes

TERMS

Physical change: A change in a physical property, for example melting, boiling.

Chemical change: How elements or compounds react with other substances.

- Physical properties do not generally depend on the amount of substance present. They do not involve a chemical reaction. Examples include density, melting point, electrical conductivity, hardness.
- A physical change involves a change in a physical property, for example melting and boiling are physical changes
- Chemical properties and chemical changes describe how elements or compounds react with other substances, for example magnesium reacts with oxygen to form magnesium oxide.

5.02 Elements

TERMS

Element: A substance containing only one type of atom.

- A chemical element contains only one type of atom in its structure.
- An element cannot be broken down into anything simpler by chemical reactions.

The structures of the elements in Figure 5.01 look different but each one of them only has one type of atom.

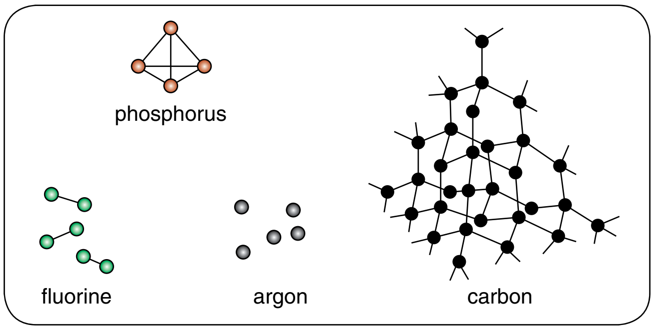


Figure 5.01

5.03 Compounds

TERMS

Compound: A substance made up of two or more different atoms (or ions) bonded together.

- When a substance contains two or more different types of atoms or ions which are bonded together it is called a compound.
- A compound has a fixed amount of each element in it. For example, a carbon dioxide molecule always has one carbon atom and two oxygen atoms.

TIP

When defining a compound, there are two things to remember: (1) there are two or more different atoms (2) the atoms are bonded (joined) together.

- Compounds usually have different properties from the elements from which they are made. The table shows how the properties of the elements sodium and chlorine differ from those of the compound sodium chloride.

Property	Sodium	Chlorine	Sodium chloride
state	solid	gas	solid
colour	silvery	green	white
boiling point / °C	883	−35	801
reaction with water	reacts to form an alkaline solution	reacts to form an acidic solution.	does not react – just dissolves

Table 5.01

5.04 Mixtures

Pure water is only made up of one substance (component). Pure sodium chloride (salt) has only one component in it.

- A **mixture** consists of two or more elements or compounds which are not chemically bonded together. For example, air is a mixture because it contains several components including oxygen, nitrogen and carbon dioxide.
- The components (parts) of a mixture can usually be separated by physical means, e.g. chromatography, distillation, crystallisation.
- A mixture can have varying numbers of each type of atom because you can change the ratio of the components. (In a compound there is always a fixed ratio of the different types of atom.)

TERMS

Mixture: An impure substance which contains two or more different components.

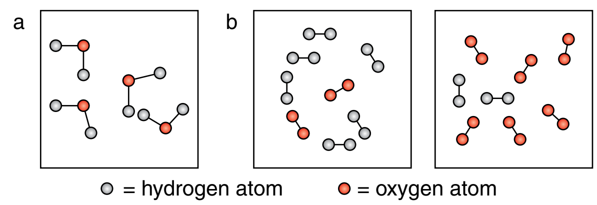


Figure 5.02 The difference between **a** a compound of oxygen and hydrogen (water) and **b** two mixtures of oxygen and hydrogen

A mixture can have varying numbers of each type of atom because you can change the ratio of the substances added together.

Sample answer

Suggest the most suitable method of obtaining purified samples of both substances in the following mixtures. In each case describe and explain why you chose the particular method.

a A solution of sodium chloride in water. [4]

Water is more volatile than sodium chloride [1] so simple distillation is used [1]. Water is the distillate [1] and sodium chloride remains in the flask [1].

b A mixture of ethanol (alcohol) and water. [3]

The two liquids have different boiling points [1], so fractional distillation is used [1]. The liquid with the lower boiling point comes off first from the distillation column [1].

c A mixture of sulfur (insoluble in water) and solid sodium chloride (soluble in water). [6]

We first need to dissolve the sodium chloride in the mixture by adding water and stirring [1] to dissolve the sodium chloride [1]. Filtration [1] is then used to filter off solid sulfur, which stays on the filter paper [1]. The solution of sodium chloride is heated to evaporate the water [1]. Crystals of sodium chloride remain in the evaporating basin [1].

Progress check

5.01 Which of the following are physical changes?

- a the reaction of nitric acid with calcium carbonate,
- b the sublimation of carbon dioxide,
- c water vapour condensing,
- d the formation of magnesium oxide when magnesium is heated in air. [2]

5.02 What is meant by the term *element*? [1]

5.03 Give two differences between a mixture and a compound. [4]

5.04 A solution of salt in water is a mixture. Explain why and describe an experiment to demonstrate your answer. [4]

5.05 Comparing compounds and mixtures

Compound (pure substance)	Mixture (impure substance)
composition has a fixed ratio of atoms	can have any composition
cannot be separated by physical means	can be separated by physical means
physical properties are different from the elements from which they are made	physical properties, e.g. colour, density, are the average of the substances in the mixture
compounds are formed by a chemical change	when mixtures are formed there is no chemical change
heat is absorbed or released when a compound is formed	heat is not usually absorbed or released when a mixture is formed (except when substances dissolve to form solutions)

Table 5.02

Worked example 5.01

Iron is a silvery metal and sulfur is a yellow non-metal. The compound iron sulfide is formed when iron and sulfur are heated together.

a Copy and complete the table to show the results of the tests. [2]

b Suggest the appearance of a mixture of iron and sulfur. [1]

c What information in the table suggests that iron sulfide is a compound and not a mixture? [2]

a

Iron and sulfur	Iron sulfide
The iron can be separated from the sulfur using a magnet.	*You need to recognise that the components in a compound cannot be separated because they are bonded together.* The iron cannot be separated from the sulfur. (1 mark)
The sulfur can be dissolved in an organic solvent leaving the iron as a solid.	Iron sulfide does not dissolve in organic solvents
You need to recognise that there is no energy change when solids are mixed. No heat is given off. (1 mark)	When the compound is formed, heat is given off.

Table 5.03

b *The appearance of a mixture is in-between that of the individual colours and in mixtures of solids, grains may be seen.* Yellow with silvery specks / silvery with yellow specks / grey [1]

c *Any piece of information from the table can be used but you must make clear that you are referring to both the compound and the mixture because both are given in the question, for example:*

The sulfur in the mixture dissolves in organic solvent [1] but the iron sulfide does not [1].

5.06 Metals and non-metals

Metals can be distinguished from non-metals by differences in their physical properties.

Physical property	Metals	Non-metals
electrical and heat conduction	conducts	do not conduct (exception – carbon as graphite).
lustre (shininess)	**lustrous**	dull surface (exceptions – iodine and graphite).
malleability (can be beaten into different shapes)	**malleable**	not malleable; brittle when hit
ductility (can be drawn into wires)	**ductile**	not ductile; break easily when a pulling force is applied
sonorous (make a ringing sound when hit)	sonorous (there are exceptions)	not sonorous; make a dull sound when hit

Table 5.04

TERMS

Ductile: Can be drawn into wires by a pulling force.

Malleable: Can be beaten into different shapes.

Lustrous: Shiny like a mirror.

Other properties which are less useful for distinguishing metals from non-metals are the following:

- Density: many metals have a high density (exceptions are the Group I metals). Most non-metals have low densities.
- Melting and boiling points: many metals have high melting and boiling points (exceptions include the Group I metals and some others, e.g. mercury, gallium). Most non-metals have low melting points (exceptions are carbon, silicon, boron).
- Hardness: many metals are hard (exceptions include the Group I metals). Most non-metals are soft (exceptions diamond (carbon), boron).

TIP

When answering questions about the physical properties of metals, it is best to select the properties that are common to all of them. For example, it would be wrong to suggest that sodium is hard and has a high melting point.

We can also use chemical properties to distinguish metals from non-metals:

- Many metallic oxides are basic. Many non-metallic oxides are acidic.
- Many metals reacts with acids to produce hydrogen. Most non-metals do not react with acids.
- When they react, metals form positive ions by losing electrons. Non-metals form negative ions by gaining electrons. (Hydrogen is an exception because it can form positive ions.)

Alloys

TERMS

Alloy: A mixture of a metal with another element or elements.

- Most alloys are mixtures of metals, e.g. brass is an alloy of copper with zinc.
- Some alloys are mixtures of metals with non-metals, e.g. mild steel is a mixture of iron with a small amount of carbon.

Progress check

5.05 Give two physical properties of most non-metals which are not shown by most metals. [2]

5.06 Give the name of two metals which are soft. [2]

5.07 Give the name of a non-metal which conducts electricity. [1]

5.08 Explain why an alloy of iron is not a pure element. [1]

Exam-style questions

Question 5.01

Sodium is a soft, shiny metal with a fairly low melting point. Bromine is a red liquid with a low boiling point. Sodium reacts with bromine to form sodium bromide, a white, crystalline solid with a high melting point.

a i Describe three properties shown by all metals. [3]

ii Give two properties of sodium that are not typical of most metals. [2]

b Give two reasons why bromine is a non-metal. [2]

c Suggest why sodium bromide is a compound of sodium and bromine and not a mixture of sodium and bromine. [4]

Question 5.02

Sulfur is a yellow non-metal that does not react with acids. Iron is a silvery, magnetic metal that reacts with hydrochloric acid to produce an odourless gas. When heated together, iron and sulfur react to produce iron sulfide. Iron sulfide is non-magnetic when pure. It is a black solid which reacts with hydrochloric acid to produce a bad-smelling gas.

a Iron sulfide is a compound. What is the meaning of the term *compound*? [1]

b Sulfur is a non-metal. Describe three properties which are typical of most non-metals. [3]

c Describe two differences between a mixture of iron and sulfur and a compound of iron and sulfur. [4]

Revision checklist

You should be able to:

- ☐ Understand the difference between physical and chemical properties
- ☐ Define the terms *element* and *compound*
- ☐ Distinguish compounds from mixtures in terms of their composition and properties
- ☐ Distinguish metals from non-metals in terms of conductivity, malleability, ductility and lustre
- ☐ Describe an alloy, such as brass, as a mixture of a metal with other elements

Unit 6

Bonding and structure

Learning outcomes

By the end of this unit you should:

- Describe the formation of ions by electron loss or gain
- Describe the formation of ionic bonds between elements from Groups I and VII
- Describe the formation of ionic bonds between metallic and non-metallic elements
- Know that ionic compounds form lattice structures containing positive and negative ions
- Describe the formation of single covalent bonds in H_2, Cl_2, H_2O, CH_4, NH_3 and HCl as the sharing of pairs of electrons leading to the noble gas configuration
- Be able to describe the electron arrangement in more complex covalent molecules
- Know how ionic and simple covalent compounds differ in volatility, solubility and electrical conductivity
- Be able to explain why ionic and simple molecular compounds differ in their melting and boiling points by referring to the forces between their particles
- Be able to describe the giant covalent structures of diamond and graphite and relate their structures to their uses
- Be able to describe the structure of silicon(IV) oxide (silicon dioxide) and relate the similarity of its properties to those of diamond
- Be able to describe metallic bonding and explain why metals are malleable and conduct electricity

6.01 The formation of ions

TERMS

Ion: An electrically charged particle formed from an atom or group of atoms by loss or gain of electrons.

- Positive **ions** are formed by the loss of one or more electrons. For example:

Na	→	Na^+	+	e^-
sodium atom		sodium ion		electron

- Negative ions are formed by the gain of one or more electrons. For example:

Cl	+	e^-	→	Cl^-
chlorine atom		electron		chloride ion

- The charge on the ion depends on the number of electrons lost or gained. For most metal ions the number of positive charges is the same as the group number. So aluminium in Group III forms 3+ ions:

$$Al \rightarrow Al^{3+} + 3e^-$$

- For non-metal ions, the negative charge is eight minus the group number. So oxygen in Group VI forms $8 - 6 = 2^-$ ions (oxide ions):

$$O + 2e^- \rightarrow O^{2-}$$

- Ions are formed by the transfer of electrons from one atom to another. We can show this by a dot-and-cross diagram (Figure 6.01).
- The ions formed in electron transfer have the electronic structure of the nearest noble gas. This is a stable electronic structure (see Figure 6.01).
- **Ionic bonds** are formed in a similar way between other elements from Group I and VII. Figure 6.02 shows the formation of lithium fluoride by transfer of an electron from a lithium atom to a fluorine atom. Only the outer shell electrons are shown.

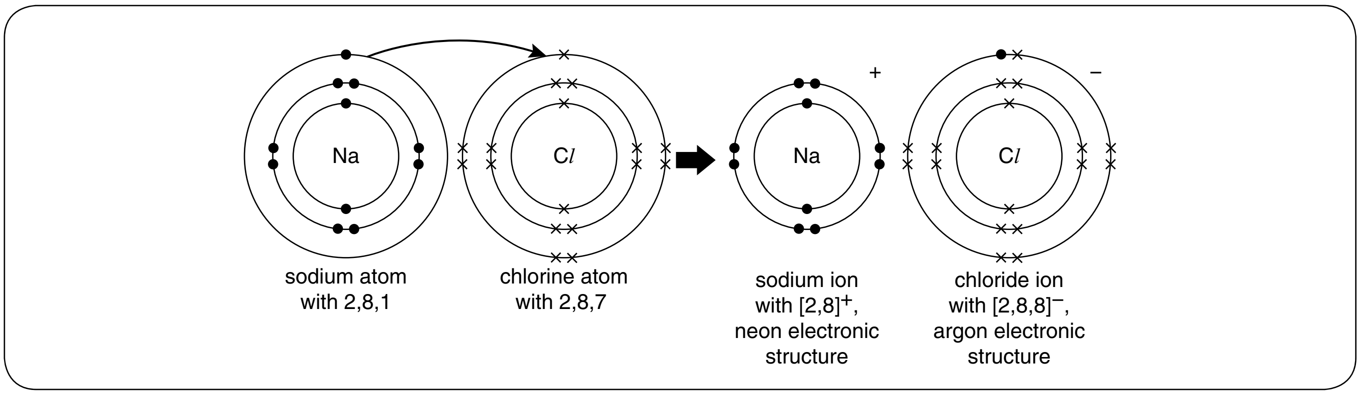


Figure 6.01

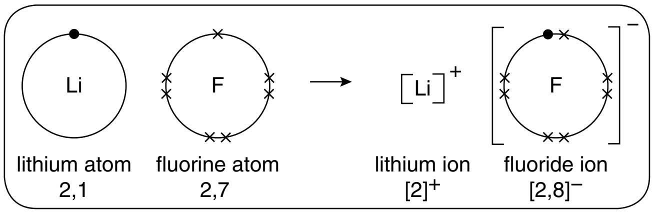


Figure 6.02

TERMS

Ionic bond: An ionic bond is a strong bond formed by the electrostatic attraction between positive and negative ions in an ionic structure.

6.02 More ionic structures

- Metals form positive ions and non-metals form negative ions.
- Ionic bonds are formed between reactive metals and reactive non-metals.

Example 1: Magnesium oxide

The two electrons in the outer shell of the magnesium atom are transferred to the oxygen atom.

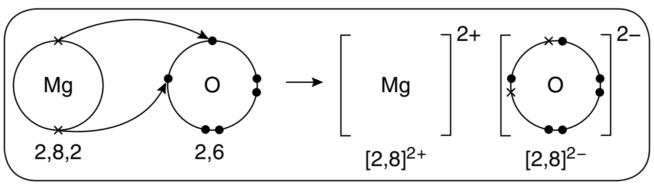


Figure 6.03

Example 2: Calcium chloride

A calcium atom loses its two outer electrons. A chlorine atom has only space in its outer shell for one electron. So, two chlorine atoms are needed to react with one calcium atom. Each chlorine atom gains one electron.

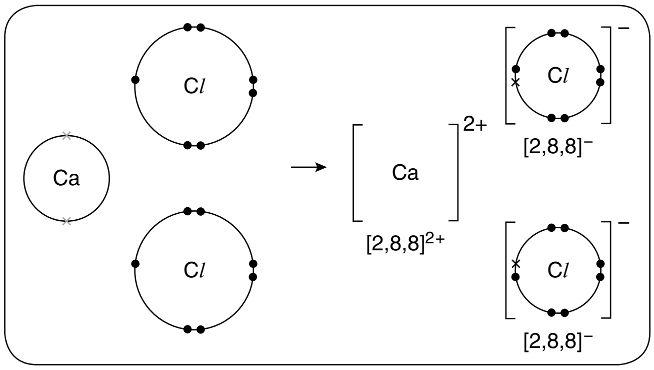


Figure 6.04

6.03 Ionic lattices

TERMS

Lattice: A continuous regular arrangement of particles which repeats itself throughout the structure.

- An ionic **lattice** has a lattice of positive and negative ions.
- In sodium chloride the ions are arranged in an alternating pattern (Figure 6.05).
- The ionic lattice is held together by strong electrostatic attractions.

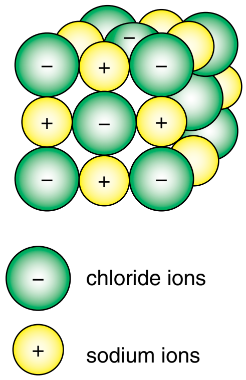


Figure 6.05

6.04 Covalent bonding

TERMS

Covalent bond: A bond formed by sharing a pair of electrons between two atoms.

Molecule: A particle having two or more atoms joined by covalent bonds.

A **covalent bond** between two hydrogen atoms in a hydrogen **molecule** is shown in Figure 6.06.

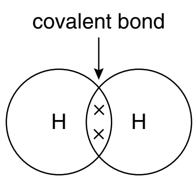


Figure 6.06

- A single covalent bond is shown by a line between the two atoms, e.g. H–H.
- The shared electrons in the covalent bond usually come from the outer shell of the atoms which combine.
- When some non-metal atoms combine, not all the electrons in the outer shell form covalent bonds. The pairs of electrons not used in covalent bonding are called **lone pairs**.

TERMS

Lone pair: A pair of electrons in the outer shell of an atom in a molecule which does not form a covalent bond.

- Some atoms are able to share two pairs of electrons to form a double bond. We show this by a double line. For example, O=O.

6.05 Drawing dot-and-cross diagrams for simple molecules

To draw a dot-and-cross diagram for a molecule:

- We usually show only the outer shell electrons.
- We use a dot for electrons from one of the atoms and a cross for the electrons of the other.
- The electrons are arranged so that the number of outer shell electrons in each atom corresponds to the nearest noble gas electronic structure.
- It is often helpful to show electrons in pairs. In the outer shell of a noble gas there are four pairs of electrons (or one pair for helium). Figure 6.07 shows how to pair the electrons in some simple molecules.
- Notice that in ammonia, NH_3, one pair of outer shell electrons on the nitrogen atom does not form a covalent bond.

Progress check

6.01 a Describe the formation of bromide ions from bromine atoms and potassium ions from potassium atoms. [2]

b Copy and complete the equation for the formation of lithium ions from lithium.

...... → Li^+ + ... [2]

6.02 Draw a dot-and-cross diagram for:

a hydrogen chloride [2]

b water. [2]

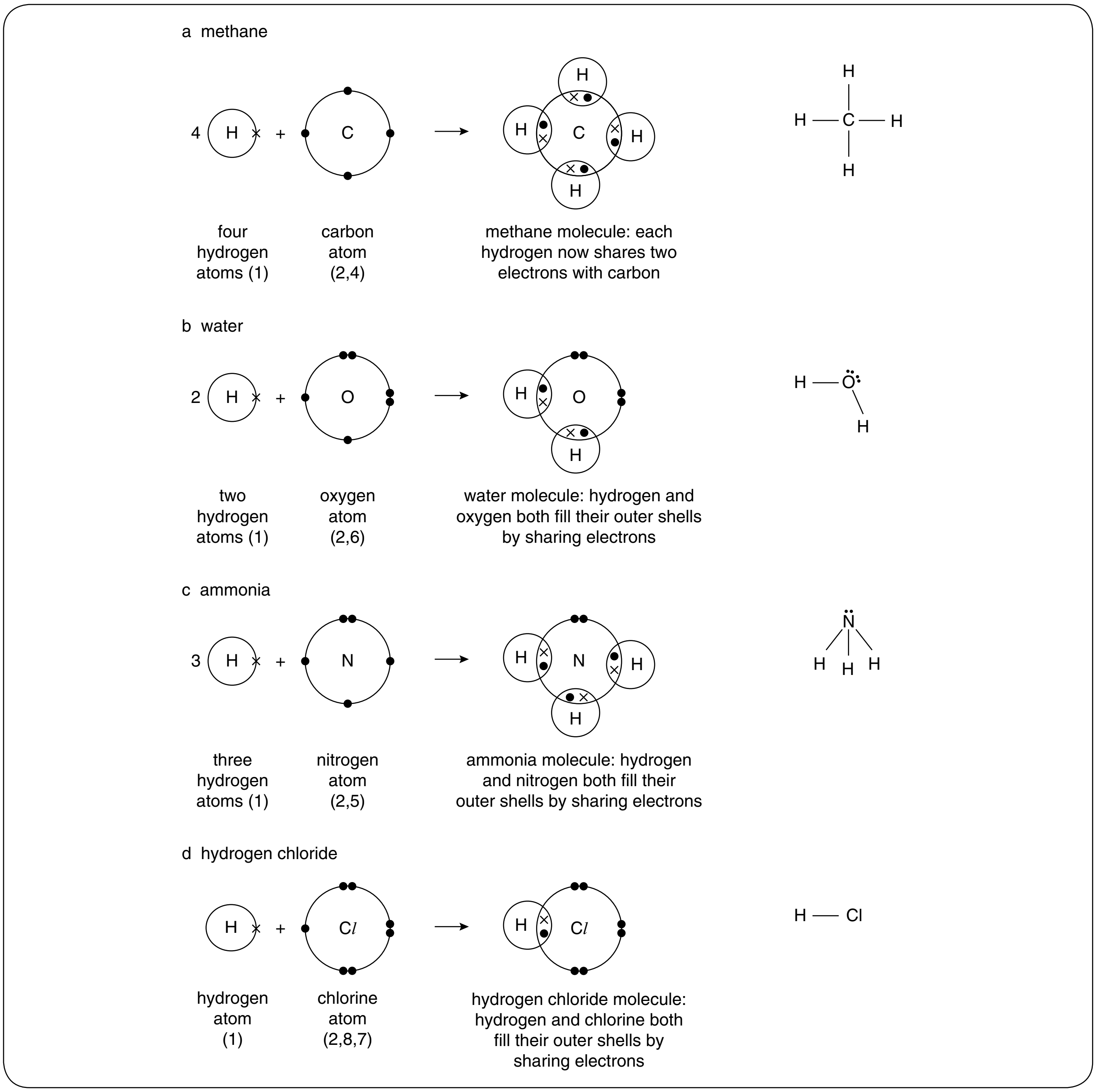


Figure 6.07

6.06 More complex covalent molecules

In an oxygen molecule, each oxygen atom (2,6) gains two electrons to complete its outer shell. It can only do this by sharing two pairs of electrons. A double bond is formed (Figure 6.08). When three pairs of electrons are shared, as in the nitrogen molecule, a triple bond is formed.

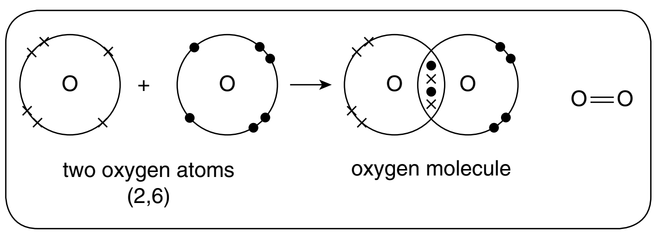


Figure 6.08

Figure 6.09 shows the dot-and-cross diagrams for some more complex molecules.

Note that if there are more than three types of atom in the molecule, you can use other symbols such as open circles or squares to show the origin of the electrons.

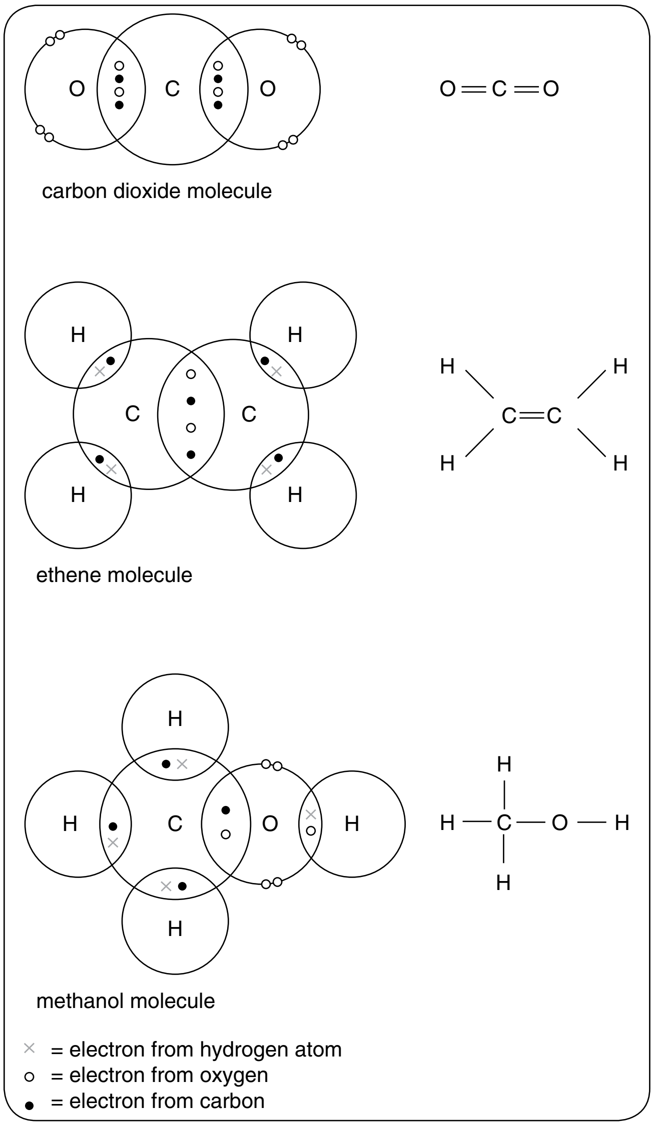


Figure 6.09

Worked example 6.01

Draw a dot-and-cross diagram to show the arrangement of the outer electrons in phosphorus trichloride, PCl_3. **[2]**

How to get the answer:

Step 1. Draw the electronic structure of a phosphorus atom and three chlorine atoms showing only the outer shell electrons (see Figure 6.10 ***a****). Refer back to Unit 4 if you are unsure how to do this.*

Step 2. Arrange the chlorine atoms so that they are around the central phosphorus atom (see Figure 6.10 ***b****).*

Step 3. Pair up the electrons, one from each chlorine atom and one from the phosphorus so that each atom has a noble gas structure. In this case it is four pairs of electrons (eight electrons) around each atom (see Figure 6.10 ***c****).*

Step 4. Make sure that there are eight electrons around the phosphorus atom by including the lone pair (see Figure 6.10 ***c****).*

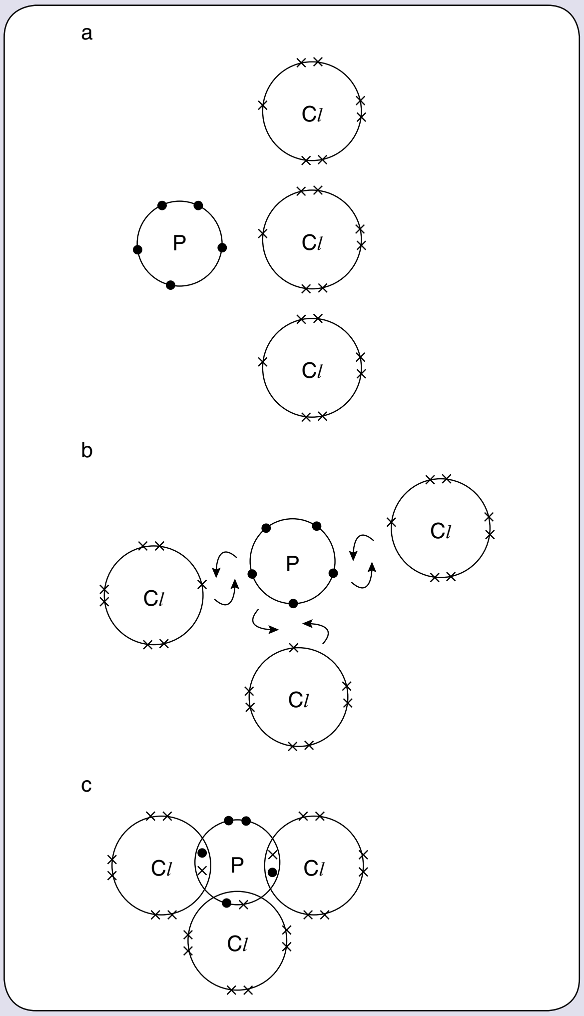


Figure 6.10

There is one mark for the correct pairing of each chlorine to the phosphorus and one mark for the lone pair on the phosphorus.

6.07 Differences between ionic and simple covalent compounds

- The particles present in ionic compounds are positive and negative ions.
- Examples of ionic compounds are salts, e.g. sodium sulfate, ammonium chloride and some metal oxides, and hydroxides, e.g. calcium oxide, sodium hydroxide.
- The particles present in simple covalent compounds are molecules.
- Examples of simple covalent molecules are octane and carbon dioxide.
- We can distinguish between ionic and simple molecular compounds by considering their volatility, solubility and electrical conductivity.

Property	Ionic compound	Simple covalent compound
volatility	Not volatile. They have high boiling points. They also have high melting points.	Many are volatile. Most have low boiling points. They also have low melting points.
solubility	Most are soluble in water but insoluble in organic solvents such as cyclohexane.	Most are insoluble in water but soluble in organic solvents. There are some exceptions.
electrical conductivity	The solids do not conduct. When molten or dissolved in water they do conduct.	Do not conduct in the solid or liquid state or in aqueous solution. There are several exceptions.

Table 6.01

- Some simple covalent compounds, e.g. ethanol, are soluble in water and not in organic solvents.
- Some simple covalent compounds, e.g. hydrogen chloride, react with water and conduct electricity in aqueous solution. This is because they form ions.

TIP

Remember that non-metallic elements having molecules, for example iodine and sulfur are simple covalent structures. These elements are volatile, insoluble in water and do not conduct electricity.

Worked example 6.02

Sodium chloride is an ionic compound.

a Describe how ions are formed from sodium and chlorine atoms. **[2]**

b Describe the electrical conductivity of solid and molten sodium chloride as well as the electrical conductivity of sodium metal and chlorine gas. **[4]**

How to get the answer:

a *Sodium is a metal. Metal ions have a positive charge. The charge on the sodium is positive because the uncharged sodium atom has lost an electron. So you just need to write:* Sodium ion is formed by loss of an electron **[1]**.

Chlorine is a non-metal. Non-metal ions have a negative charge. The charge on the chloride ion is positive because the uncharged chlorine atom has gained an electron. So you just need to write: A chloride ion is formed by the gain of an electron **[1]**.

b *Sodium chloride is an ionic solid. For a substance to conduct electricity it needs moving ions or electrons flowing through the whole structure. The ions in solid sodium chloride cannot move because they are in fixed positions. So 1 mark is awarded for* 'solid sodium chloride does not conduct electricity' **[1]**.

When ionic compounds are molten (liquid) the ions are free to move. So 1 mark would be awarded for 'molten sodium chloride conducts electricity' **[1]**.

Sodium is a metal. All metals conduct electricity. So 1 mark is awarded for 'sodium conducts electricity' **[1]**.

Chlorine is a non-metal and is a simple molecular compound. Non-metals (except graphite) do not conduct electricity. So 1 mark would be awarded for 'chlorine does not conduct electricity' **[1]**.

Progress check

6.03 A compound which is soluble in water has a melting point of 801 °C. What type of structure does this substance have? [1]

6.04 Compound X has a simple molecular structure. Suggest the properties of compound X in terms of volatility, electrical conductivity and solubility in water. [3]

6.08 Explaining differences in properties

- Ionic compounds have high melting and boiling points because of the strong attractive forces between the ions. It takes a lot of energy to overcome these forces.
- Simple covalent compounds have low melting and boiling points because the forces between their molecules (**intermolecular forces**) are weak.

TERMS

Intermolecular forces: The weak forces between molecules.

- Ionic compounds do not conduct electricity when in the solid state. This is because the ions cannot move. They do conduct in the molten (liquid) state or when dissolved in water because the ions separate out and can move.
- Simple covalent compounds do not conduct electricity when solid or molten. This is because they do not have ions or electrons to conduct. Molecules are not electrically charged.

TIP

Make sure that you use the phrase 'intermolecular forces' correctly. The forces between molecules are weak but the forces between the atoms within the molecule are strong covalent bonds.

Sample answer

a Sodium oxide, Na_2O is an ionic compound with a lattice structure.

Draw a dot-and-cross diagram for sodium oxide showing only the outer electron shells. Include the correct ionic charges and the correct number of ions. [4]

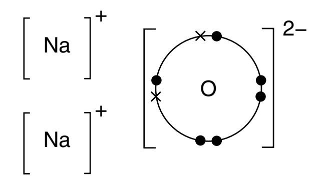


Fig 6.11

1 mark for showing 2 sodium ions and 1 oxide ion.

1 mark for the correct electronic configuration in each ion.

1 mark for each correct charge Na^+ **[1]** and O^{2-} **[1]**.

b i What is meant by the term ionic lattice? [3]

A regularly repeating arrangement **[1]** of positive and negative ions **[1]** which repeats itself throughout the structure **[1]**.

ii Explain why sodium oxide has a high melting point and does not conduct electricity at room temperature. [4]

Sodium oxide has a high melting point because it has strong forces between the ions **[1]** because the positive and negative ions attract each other electrostatically **[1]**. It needs a lot of energy to overcome these strong forces **[1]**.

It does not conduct because the ions are not free to move **[1]** in the solid.

6.09 Giant covalent structures

TERMS

Giant covalent structures (macromolecular structures): Structures with a lattice (network) of covalent bonds which repeats throughout the whole structure.

Diamond and graphite are both **giant covalent structures** made of carbon atoms (Figure 6.12).

In diamond, each carbon atom forms four covalent bonds with other carbon atoms. This network of strong bonds extends unbroken throughout the whole structure. In graphite, the carbon atoms are hexagonally arranged in layers.

TIP

Make sure that you do *not* use the term intermolecular forces when referring to diamond. If referring to graphite, only use the term intermolecular forces when writing about the forces *between* the layers.

Similar properties of diamond and graphite

The properties of diamond and graphite can be explained by their structure and bonding:

- They have high melting and boiling points. It is difficult to break down the network of strong covalent bonds. It needs a high temperature to overcome the huge numbers of strong bonds and melt the solid.
- They are insoluble in water and in organic solvents. The network of covalent bonds is too strong to allow solvent molecules to form strong enough bonds with the carbon atoms.

TIP

When answering questions about why giant structures have high melting points it is incorrect to write about strong forces between *molecules*. The best answers refer to strong bonds between the *atoms*.

How the properties of diamond and graphite differ

Hardness:

- Diamond is hard. The network of strong covalent bonds makes it difficult to scratch the surface of the crystal. Because it is so hard, diamond is used for the edges of tools such as glass cutters and drill bits.
- Graphite is soft. It is easy to scratch. The forces between the layers of graphite are weak. The layers can slide over each other when a force is applied. The layers flake away easily. So graphite is used as a lubricant and in the 'leads' of pencils.

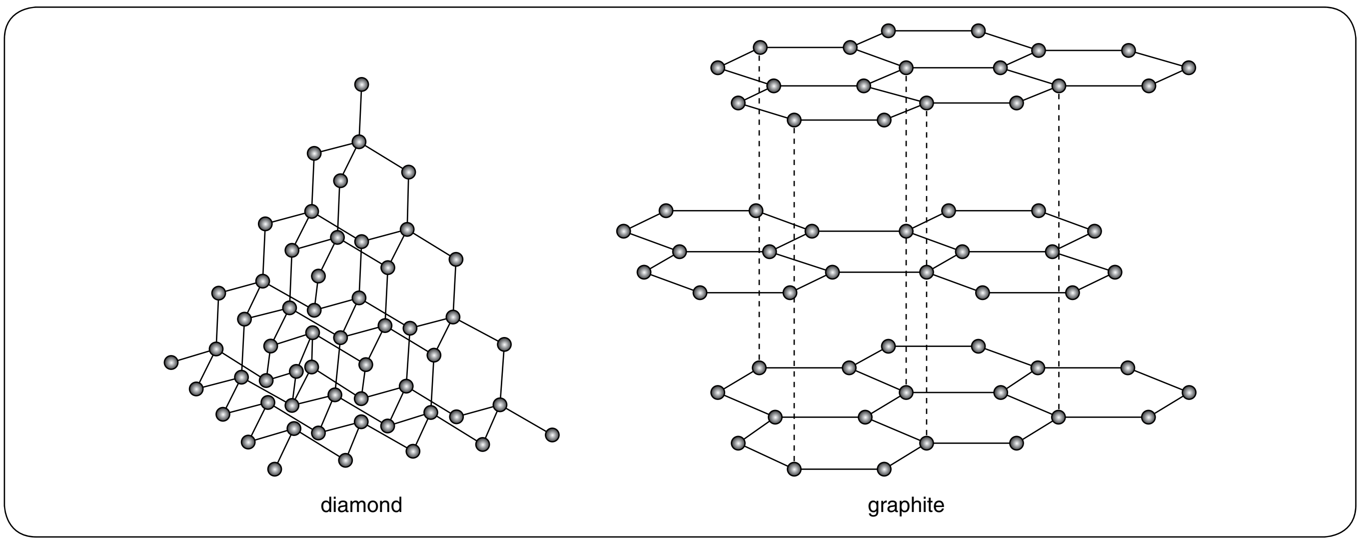


Figure 6.12