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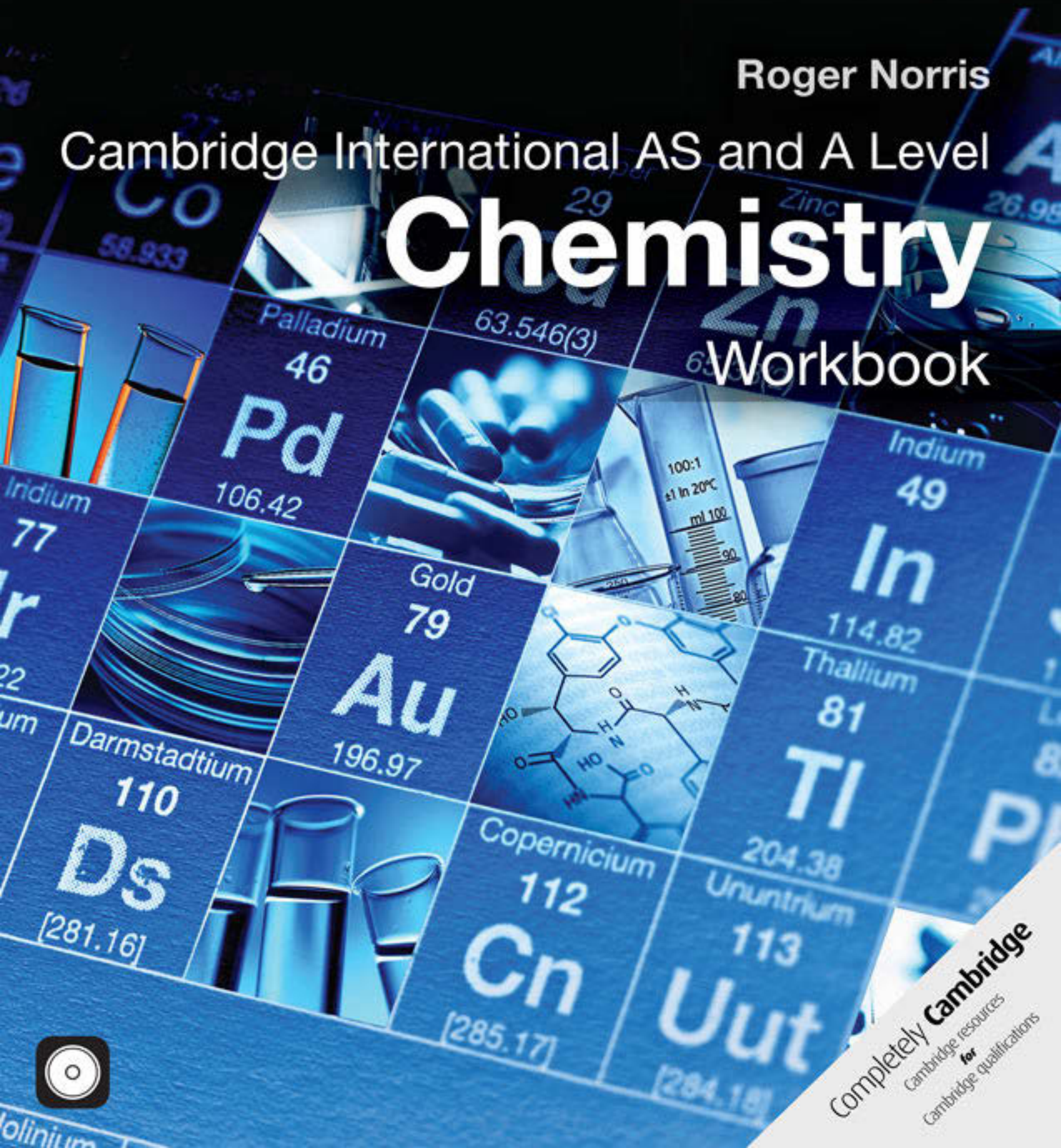Endorsed for learner support

Roger Norris

Cambridge International AS and A Level
Chemistry
Workbook

Completely **Cambridge** — Cambridge resources for Cambridge qualifications

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Cambridge International AS and A Level

Chemistry

Workbook

How to use this book

A **Chapter outline** appears at the start of every chapter to set the scene and to help with navigation through the book

Chapter outline

■ Relative atomic mass, isotopic mass and formula mass based on the ^{12}C scale
■ Empirical formula and molecular formula
■ Using mass spectra to calculate relative atomic mass
■ Constructing balanced equations
■ Performing calculations involving moles, reacting masses, gas volumes and solution concentration

A list of **Key terms** are defined and explained clearly at the start of each topic

KEY TERMS

Atomic orbitals: Regions of space outside the nucleus occupied by a maximum of two electrons. The s, p, d and f orbitals have different shapes.

Electronic configuration: Way of showing the arrangement of electrons in an atom, e.g. $1s^2 2s^2 2p^5$.

Energy levels (of electrons): Regions at various distances from the nucleus where electrons have particular amounts of energy. The further the energy level is from the nucleus, the greater the amount of energy it has.

Ionisation energy, ΔH_{i1}: The first ionisation energy is the energy needed to remove 1 mole of electrons from 1 mole of atoms of an element in the gaseous state to form 1 mole of gaseous ions. Successive ionisation energies refer to further electrons being removed, $\Delta H_{i2}, \Delta H_{i3}$, etc.

Principal quantum shells: The main energy levels in the atoms where electrons have particular amounts of energy. The first quantum shell can hold a maximum of 2 electrons, the second 8 and the third 18.

Shielding: The ability of inner shell electrons to reduce the effect of the nuclear charge on outer shell electrons.

Subshells: Regions of the principal quantum shells where electrons exist in defined areas associated with particular amounts of energy.

Each **Exercise** in every chapter helps students to practise the necessary skills for studying Chemistry at AS and A Level

Exercise 3.1 Electron shells and subshells

This exercise is designed to support your understanding of electron shells and subshells. It also gives practice at deducing and interpreting electron configurations.

a The table gives information about electron shells and atomic orbitals. Deduce the numbers and types of **atomic orbitals** represented by the letters A to F.

principal quantum shell	maximum number of electrons in principal quantum shell	number of subshells in principal quantum shell	types of orbital present
1	A	1	B
2	8	C	2s, 2p
3	D	E	F

Detailed **Learning support** is provided throughout to help students tackle the different exercises and build confidence to answer questions independently.

Part (b) asks you to interpret graphs. Remember:

- Rate is proportional to $\dfrac{\text{change in amount of product made (or reactant used up)}}{\text{time}}$.

- Read the stem of the question carefully to get important pieces of information.

Exam-style questions allow students to thoroughly prepare for examinations and check answers which are provided on the CD-ROM at the back of the book.

Exam-style questions

QUESTION 1

Hydrogen peroxide decomposes slowly at r.t.p. The reaction is exothermic.

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$
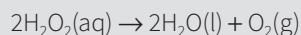
a How is the rate of reaction affected by each of the following conditions? In each case give an explanation in terms of collision theory or activation energy.

i Decreasing the temperature. [2]

ii Increasing the pressure. [2]

iii Adding a catalyst. [2]

Contents

Chapter 1
Moles and equations

Chapter outline

- Relative atomic mass, isotopic mass and formula mass based on the ^{12}C scale
- Empirical formula and molecular formula
- Using mass spectra to calculate relative atomic mass
- Constructing balanced equations
- Performing calculations involving moles, reacting masses, gas volumes and solution concentration

KEY TERMS

Avogadro constant, _L_: The number of defined particles (atoms, ions, molecules or electrons) in a mole of those particles. Its value is 6.02×10^{23}.

Empirical formula: The simplest ratio of the different atoms present in a molecule.

Mass spectrometer: An instrument for helping to deduce the relative isotopic abundance of elements and the structure of compounds.

Molar gas volume: The volume of one mole of gas at a defined temperature and pressure. At **r.t.p** its value is $24.0\,dm^3$.

Molar mass: The mass of a mole of substance in grams.

Mole: The amount of substance that has the same number of particles (atoms, ions, molecules or electrons) as there are atoms in exactly $12\,g$ of the ^{12}C isotope.

Molecular formula: The actual number of each type of atoms in a molecule.

Relative atomic mass: The weighted average mass of the atoms of an element on a scale on which an atom of the ^{12}C isotope has a mass of exactly 12 units.

Relative formula mass: The weighted average mass of one formula unit of a compound measured on a scale on which an atom of the ^{12}C isotope has a mass of exactly 12 units.

Relative isotopic mass: The mass of a particular isotope of an element on a scale on which an atom of the ^{12}C isotope has a mass of exactly 12 units.

Relative molecular mass: The weighted average mass of a molecule on a scale on which an atom of the ^{12}C isotope has a mass of exactly 12 units.

r.t.p.: Room temperature and pressure (1 atmosphere and $20\,°C$).

Spectator ion: Ion present in a reaction mixture which does not take part in the reaction.

State symbol: Symbol placed after each reactant and product to indicate whether they are solid (s), liquid (l), gas (g) or in aqueous solution (aq).

Stoichiometry: The mole ratio of the reactants and products in a chemical equation.

Exercise 1.1 Definitions

This exercise will familiarise you with some important definitions about relative masses, moles and the Avogadro constant.

It is important that you learn definitions carefully. Remember that:

- A mole is an amount of substance.
- You compare this amount with an atom of the ^{12}C isotope.
- Relative atomic masses are weighted averages.

Copy and complete these sentences using words from the list

adding   amount   atomic   atoms   Avogadro   carbon-12   constant

grams   ionic   isotope   mole   relative   twelve   weighted

A **mole** is the _____ of substance that has the same number of defined particles as there are _____ in exactly _____ grams of the _____ isotope. This number of particles is called the _____ _____ .
Relative atomic mass is the _____ average mass of atoms of an element on the scale where an atom of the ^{12}C _____ has a mass of 12 units. **Relative molecular mass** is found by _____ together the _____ _____ masses of the atoms in a compound. For _____ compounds we use the term **relative formula mass**. Another term is **molar mass** which is the mass of a _____ of substance in _____ .

Exercise 1.2   Mole calculations

This exercise will familiarise you with some basic calculations using the mole concept.

- $\text{moles} = \dfrac{\text{mass (in g)}}{\text{molar mass (in mol dm}^{-3})}$.
- In part (a) remember to use the mole ratio.
- In part (a)(iv) you need to rearrange the equation: (mass (in g) = moles × molar mass (in mol dm⁻³).

a   Lead oxide, Pb_3O_4, is reduced by heating with excess carbon.

$$Pb_3O_4 + 4C \rightarrow 3Pb + 4CO$$
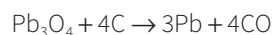
Use the following method to calculate the maximum mass of lead formed when 41.12 g of Pb_3O_4 is reduced.
Calculate:

i   The molar mass of Pb_3O_4 (A_r values: Pb = 207.2, O = 16.0).

- The number of significant figures in your answer should be the same as the least number of significant figures in the data.
- 326.7 is to 4 significant figures.
- 0.014 is to 2 significant figures.

ii   The amount in moles of Pb_3O_4 (to 3 significant figures).

iii   The amount in moles of lead produced.

iv   Mass of lead produced (to 3 significant figures).

b   35.61 g of tin, reacts with exactly 42.60 g of chlorine, Cl_2, to form 78.21 g of tin(IV) chloride, $SnCl_4$.

i   Calculate the number of moles of tin, chlorine and tin chloride. (A_r values: Sn = 118.7, Cl = 35.5)

ii   Deduce the **stoichiometry** of the reaction.

iii   Write a balanced equation for the reaction.

Exercise 1.3 Deducing formulae and composition by mass

This exercise will help you deduce **empirical formula** and **molecular formula** as well as percentage composition by mass.

Remember that % by mass $= \dfrac{\text{atomic mass} \times \text{number of moles of element}}{\text{molar mass of compound}}$

a When 14.98 g of arsenic are completely combusted, 19.78 g of an oxide of arsenic are formed. Calculate

  i   The mass of oxygen in this oxide of arsenic.

  ii The amount in moles of atoms of arsenic and oxygen which combine. (A_r values: As = 74.9, O = 16.0)

  iii The empirical formula.

b The molar mass of this oxide of arsenic is 395.6 g. Deduce the molecular formula of this oxide of arsenic.

c The empirical formula of another oxide of arsenic is As_2O_5.

  Calculate the percentage by mass of arsenic in As_2O_5. Give your answer to 3 significant figures.

Exercise 1.4 Using molar gas volume

This exercise helps you use the **molar gas volume** to deduce the stoichiometry of a reaction.

Remember that:

* The volume of one mole of gas at **r.t.p.** is 24 dm³.

* Moles of gas $= \dfrac{\text{volume (in dm}^3)}{24}$ or $\dfrac{\text{volume (in cm}^3)}{24\,000}$.

a Deduce the volume, number of moles or mass of gas represented by the letters A to F. (A_r values: P = 31.0, O = 16.0, S = 32.1, H = 1.00)

gas	volume of gas	moles of gas / mol	mass of gas / g
PH_3	80.0 cm³	A	B
SO_2	C dm³	D	8.00 g
O_2	E cm³	0.150 mol	F

b Two syringes are set up as shown.

A   Tap   B

40 cm³ oxygen, O_2 40 cm³ nitrogen(II) oxide, NO
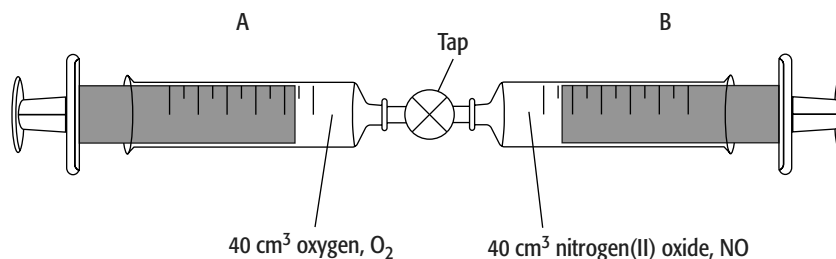
Small measured volumes of oxygen were pushed from syringe A into syringe B. The product is another gaseous oxide of nitrogen, NO_y. After each addition of oxygen, the tap was closed and the total volume of gases measured. The results are shown below.
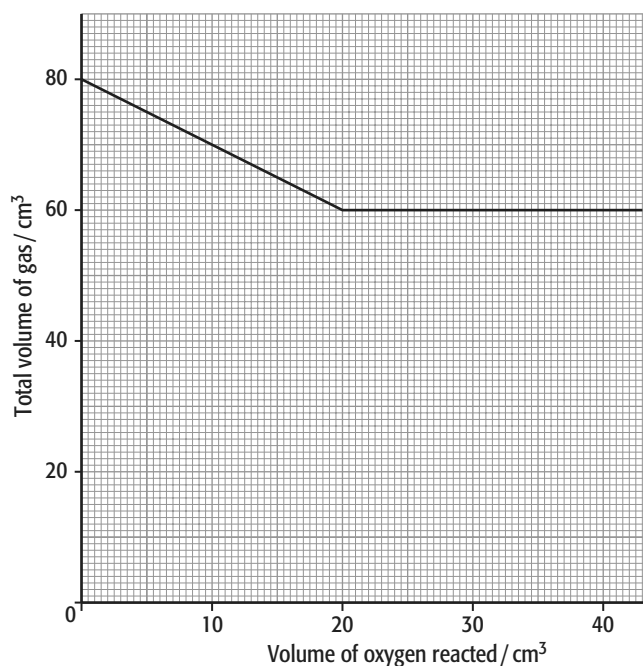
i   What volume of oxygen reacts with 40 cm³ of nitrogen(II) oxide?

ii   What volume of NO_y is formed?

iii   Deduce the formula of NO_y

iv   Write a balanced equation for the reaction.

Exercise 1.5   Solution concentration

This exercise gives you practice in calculating volumes, moles and concentrations. It also revises calculations from titration results.

Remember that:

*   Concentration (in mol dm⁻³) $= \dfrac{\text{amount (in mol)}}{\text{volume (in dm}^3\text{)}}$.

*   In some questions you will have to rearrange this equation.

a   Deduce the values represented by the letters R to V. (A_r values: Na = 23.0, O = 16.0, Cl = 35.5, H = 1.0)

solute	moles or mass of solute	volume of solution	concentration of solution
$CuSO_4$	0.12 mol	200 cm³	R
HCl	S mol	1.5 dm³	0.4 mol dm⁻³
$ZnCl_2$	0.25 mol	T cm³	0.05 mol dm⁻³
NaOH	5.4 g	150 cm³	U
NaCl	V g	0.20 dm³	2.0 mol dm⁻³

b 20.0 cm³ of a solution of barium hydroxide, $Ba(OH)_2$, is exactly neutralised by 35.4 cm³ of 0.200 mol dm⁻³ hydrochloric acid.

$$Ba(OH)_2 + 2HCl \rightarrow BaCl_2 + 2H_2O$$

Calculate

i The amount in moles of HCl.

ii The amount in moles of $Ba(OH)_2$.

iii The concentration of $Ba(OH)_2$.

Express your answers to 3 significant figures.

Exercise 1.6 Writing equations

This exercise provides practice in balancing equations, including ionic equations, as well as the use of state symbols.

Remember that when writing equations:

- You must not alter the formula of a compound.
- The number of atoms of each type must be the same on each side of the equation.
- Balance only by changing the numbers in front of particular compounds.

a Aqueous barium nitrate, $Ba(NO_3)_2$, reacts with dilute sodium sulfate, Na_2SO_4. A precipitate of barium sulfate is formed as well as one other aqueous compound.

i Write a balanced equation for this reaction. Include state symbols.

ii Convert the equation in part **i** into an ionic equation.

iii Name the **spectator ions** in this reaction.

b 2.50×10^{-2} mol of dilute hydrochloric acid reacts exactly with 1.25×10^{-2} mol of an insoluble oxide of iron. The products are aqueous iron(II) chloride and water.

Deduce the balanced equation for this reaction. Include state symbols.

c Convert the word equation below into a balanced ionic equation. Include state symbols for the substances at r.t.p. (calcium chloride is soluble in water).

calcium carbonate + hydrochloric acid → calcium chloride + carbon dioxide + water
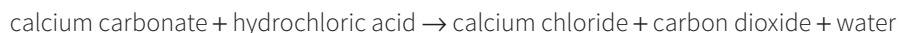
Exercise 1.7 Accurate relative molecular masses

The mass spectrometer gives information about the number and mass of the isotopes present in a sample of an element. In this exercise you will be using relative isotopic masses to calculate an accurate value for the relative atomic mass of strontium.

When calculating accurate relative atomic masses:

- Multiply each isotopic mass by its % abundance.
- Add these figures and divide by 100.

The mass spectrum shows the relative abundance of the isotopes present in a sample of strontium.
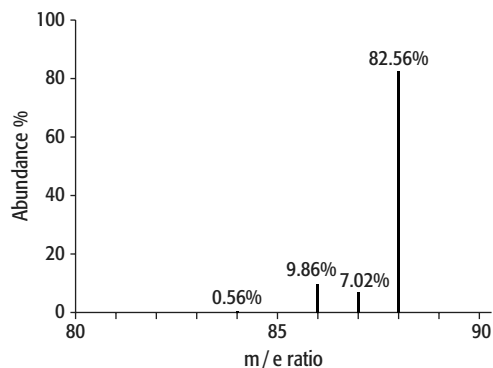
a Write the symbol for the lightest isotope present.

b Use the data in the mass spectrum to calculate the relative atomic mass of strontium. Express your answer to 3 significant figures.

- In part (a) the term *describe* means that you have to give details of the procedure.
- In part (b) remember to take the mole ratio of acid to alkali into account.
- When doing calculations make sure that you show all your working.
- Remember not to round up to the correct number of significant figures until the end of the calculation.

Exam-style questions

QUESTION 1

A sample of 3.60 g of malic acid, $C_2H_4O(CO_2H)_2$, was dissolved in 20.0 cm³ of distilled water. The solution was titrated with 0.125 mol dm⁻³ aqueous sodium hydroxide.

a Describe how you would carry out this titration. [4]

b The equation for the reaction is

$$C_2H_4O(COOH)_2 + 2NaOH \rightarrow C_2H_4O(COONa)_2 + 2H_2O$$
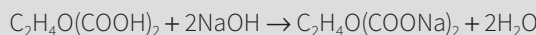
Calculate the volume of aqueous sodium hydroxide used. Express your answer to 3 significant figures. [4]

c 25 cm³ of a 0.0125 mol dm⁻³ solution of a metal hydroxide, $X(OH)_y$ was titrated with 0.05 mol dm⁻³ hydrochloric acid. It required 12.5 cm³ of acid to neutralise the hydroxide.

Deduce the value of y and write a balanced equation for the reaction. [4]

Total: 12 marks

QUESTION 2

- In part (a) make sure that you give an accurate definition.
- In part (b) you need to make sure that you are calculating the number of atoms and not molecules.
- In part (c) remember that equal volumes of gases have equal numbers of moles of molecules.

a What is meant by the term *Avogadro constant*? [1]

b How many oxygen atoms are there in 0.0011 g of carbon dioxide? ($L = 6.02 \times 10^{23}$, A_r values: C = 12.0, O = 16.0) [3]

c 14 cm³ of butene gas, C_xH_y, reacts with exactly 84 cm³ of oxygen. 56 cm³ of carbon dioxide is formed. Deduce the formula of butene. Show all working. [4]

d i A compound has the following percentage composition by mass: C 37.25%, H 7.75%, Cl 55%. Deduce the empirical formula. [3]

ii What further information is needed to deduce the molecular formula of this compound? [1]

Total: 12 marks

- Part (a) involves calculation of relative atomic mass.
- In calculations (b) and (c) make sure that you show all working clearly, e.g.

moles of $Fe_2O_3 = \dfrac{mass}{M_r}$

QUESTION 3

The table shows the relative abundances of the four naturally occurring isotopes of iron.

isotopic mass	relative abundance
54	5.840
56	91.680
57	2.170
58	0.310

a Calculate the relative atomic mass of iron to 3 significant figures. [3]

b Limonite is a mineral with the formula $Fe_2O_3.H_2O$. Calculate the percentage by mass of iron in limonite. [3]

c **i** Calculate the maximum mass of iron formed when 798 g of iron(III) oxide, Fe_2O_3, is reduced by excess carbon monoxide.

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

Express your answer to 3 significant figures. (A_r values Fe = 55.8, O = 16.0) [3]

ii Calculate the volume of carbon dioxide formed at r.t.p. [2]

d Red hot iron reacts with steam to form Fe_3O_4 and hydrogen. Write a balanced equation for this reaction. [2]

e Iron reacts with aqueous copper(II) sulfate. The products are copper and aqueous iron(II) sulfate. Construct the ionic equation for this reaction including state symbols. [2]

Total: 15 marks

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Chapter 2
Atomic structure

Chapter outline

- The relative charges and masses of protons, neutrons and electrons
- How protons, neutrons and electrons behave in electric fields
- Deducing the number of protons and neutrons in atoms and ions
- Describing proton number, nucleon number and isotopes
- Using the symbolism $^{x}_{y}A$ for isotopes.

KEY TERMS

Atomic number: The number of protons in the nucleus of an atom.

Electron: Negatively charged particle found in orbitals outside the nucleus of an atom. It has negligible mass compared with a proton.

Isotope: Atom of an element with the same number of protons but a different number of neutrons.

Mass number: See nucleon number.

Neutron: Uncharged particle in the nucleus of an atom, having the same relative mass as a proton.

Nucleon number (Mass number): The number of protons plus neutrons in the nucleus of an atom.

Nucleus: The dense core at the centre of an atom containing neutrons (except the ^{1}H isotope) and protons.

Proton: Positively charged particle in the nucleus of an atom.

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Exercise 2.1 Atomic structure

This exercise will familiarise you with the properties of the three types of subatomic particle.

Copy and complete these sentences using words from the list below. Some words may be used more than once.

electron   negative   neutrons   positively   protons

relative   shells

An atom contains a dense **nucleus** surrounded by _____ of **electrons**. The nucleus contains the nucleons (_____ and _____). **Protons** are _____ charged, electrons have a _____ charge and _____ are uncharged. The _____ and **neutrons** have the same _____ mass. The mass of an _____ is negligible.

Exercise 2.2 Terms used in atomic structure

This exercise will familiarise you with some terms related to isotopes.

Match the boxes 1 to 5 on the left with the descriptions A to E on the right.

1. **Atomic number**	A. Another name for nucleon number
2. **Isotopes**	B. The number of protons plus neutrons in the nucleus
3. **Mass number**	C. The number of protons in the nucleus of an atom
4. **Neutrons**	D. Atoms with the same number of protons but different numbers of neutrons
5. **Nucleon number**	E. Uncharged particles in the nucleus

Exercise 2.3 Isotopes

This exercise will familiarise you with the concept of isotopes and help you deduce the number of particular subatomic particles in an atom.

- When defining terms, you must be precise. Remember that the term *isotope* refers to atoms and not elements, molecules or substances.
- Number of neutrons = nucleon number − atomic number
- Alternative terms you should know are: mass number for nucleon number and proton number for atomic number.
- The top number in isotopic formulae is the number of protons + neutrons and the bottom number is the atomic number.

a Deduce the number of protons and electrons or neutrons represented by the letters A to F.

isotope	number of protons	number of electrons	number of neutrons
$^{86}_{36}Kr$	36	A	50
$^{115}_{49}In$	B	49	C
$^{50}_{24}Cr$	D	E	F

b A 'cell' of the Periodic Table is shown below:

38
Sr
87.6

   i Explain why the relative molecular mass is not a whole number.

   ii An isotope of strontium has a nucleon number of 90. How many neutrons are there in this isotope?

   iii Explain in terms of the charge on the subatomic particles why the strontium ion has a 2+ charge.

c How many protons, neutrons and electrons do the following species have?

   i $^{27}_{13}Al$ **ii** $^{133}_{55}Cs^+$ **iii** $^{17}_{8}O^{2-}$

Exercise 2.4 The discovery of the nucleus

This exercise explores how the nucleus was discovered and will familiarise you with the behaviour of charged particles.

You should be able to answer questions about unfamiliar material. Always:

- Read the information carefully, noting down the key points.
- Take note of the command words such as *explain* and *suggest*. *Explain* means that you have to use scientific ideas or theories to state why something happens. *Suggest* is often used instead of explain, when the material is unfamiliar or there is more than one possible answer.

In 1910, researchers in Manchester, UK, fired alpha-particles (α-particles) at thin sheets of gold foil. Some of the α-particles passed straight through the foil (course A in the figure below). Others were deflected slightly (course B). About 1 in every 20 000 was deflected backwards (course C).
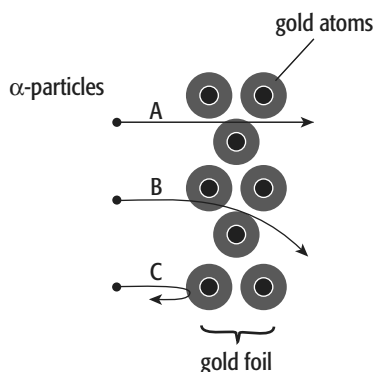
a   Alpha-particles are helium nuclei. Helium atoms have 2 protons and 2 neutrons.

Write the isotopic symbol for a helium nucleus.

b   Suggest, in terms of the structure of the atoms, why most α-particles passed straight through the foil.

c   Explain why some α-particles were deflected slightly.

d   Suggest, in terms of the structure of the atoms, why so few α-particles were deflected backwards.

e   Suggest what would happen in this experiment if a beam of neutrons were fired at the gold foil. Explain your answer.

Exam-style questions

QUESTION 1

This question is about isotopes and subatomic particles.

In part (b) you need to refer back to the Avogadro constant from Chapter 1.

The diagram shows the structure of an isotope of lithium.

a Describe the number and charge of each subatomic particle present [3]

b How many nucleons are present in one mole of this isotope of lithium? ($L = 6.02 \times 10^{23}$) [2]

c Write the isotopic symbol for a lithium ion. [2]

d Explain why a lithium ion is positively charged. [1]

Total: 8 marks

When answering questions it is important to note the number of marks available. For example in parts (b) and (d) you need to give at least three separate points in order to gain full marks.

QUESTION 2

Cobalt and nickel are next to each other in the Periodic Table.

27	28
Co	Ni
58.9	58.7

a Which one of these elements has the higher atomic number? Explain your answer. [1]

b Explain why nickel has a lower relative atomic mass than cobalt. [3]

c The formulae of two isotopes are:

$^{59}_{27}$Co $^{58}_{28}$Ni

   i Which one of these isotopes has a greater number of neutrons? Explain your answer. [1]
   ii Which one of these isotopes has fewer electrons? Explain your answer. [1]
   iii An ion of cobalt has 27 protons and 24 electrons. Give the formula for this ion. [1]

d A beam of electrons is fired through an electric field between two charged plates.

beam of electrons
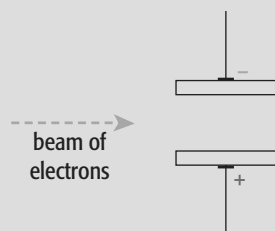
Describe how the electron beam behaves when it passes through the plates. Explain your answer. [3]

Total: 10 marks

Chapter 3
Electrons in atoms

Chapter outline

- The number and relative energies of s, p and d orbitals for quantum shells 1, 2 and 3 and the 4s and 4p orbitals
- The shapes of s and p orbitals
- Electronic configuration of atoms and ions, e.g. $1s^22s^22p^6$, etc.
- Ionisation energy and the factors influencing it
- Ionisation energy data used to explain trends in periods and groups in the Periodic Table
- Interpreting successive ionisation energy data

> **KEY TERMS**
>
> **Atomic orbitals:** Regions of space outside the nucleus occupied by a maximum of two electrons. The s, p, d and f orbitals have different shapes.
>
> **Electronic configuration:** Way of showing the arrangement of electrons in an atom, e.g. $1s^22s^22p^5$.
>
> **Energy levels (of electrons):** Regions at various distances from the nucleus where electrons have particular amounts of energy. The further the energy level is from the nucleus, the greater the amount of energy it has.
>
> **Ionisation energy, ΔH_{i1}:** The first ionisation energy is the energy needed to remove 1 mole of electrons from 1 mole of atoms of an element in the gaseous state to form 1 mole of gaseous ions. Successive ionisation energies refer to further electrons being removed, $\Delta H_{i2}, \Delta H_{i3}$, etc.
>
> **Principal quantum shells:** The main energy levels in the atoms where electrons have particular amounts of energy. The first quantum shell can hold a maximum of 2 electrons, the second 8 and the third 18.
>
> **Shielding:** The ability of inner shell electrons to reduce the effect of the nuclear charge on outer shell electrons.
>
> **Subshells:** Regions of the principal quantum shells where electrons exist in defined areas associated with particular amounts of energy.

Exercise 3.1 Electron shells and subshells

This exercise is designed to support your understanding of electron shells and subshells. It also gives practice at deducing and interpreting electron configurations.

a The table gives information about electron shells and atomic orbitals. Deduce the numbers and types of **atomic orbitals** represented by the letters A to F.

principal quantum shell	maximum number of electrons in principal quantum shell	number of subshells in principal quantum shell	types of orbital present
1	A	1	B
2	8	C	2s, 2p
3	D	E	F

- Make sure that you learn the order of the filling of the electrons in subsidiary quantum shells, especially from the 3rd energy level onwards.
- For Groups 1 and 2, the number of electrons in the outer shell of an uncharged atom is equal to the group number.

b The table shows the electron configuration of some atoms and ions. Deduce the missing information represented by the numerals (i) to (vii).

proton number	symbol	electronic configuration
9	F	(i)
14	(ii)	$1s^2 2s^2 2p^6 3s^2 3p^2$
24	Cr	(iii)
11	Na^+	(iv)
(v)	K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
35	(vi)	$[Ar]3d^{10}4s^2 4p^6$
22	Ti^{2+}	(vii)

c To which group do the atoms with these electron configurations belong?

i $1s^2 2s^2 2p^6 3s^2$

ii $1s^2 2s^2 2p^3$

iii $1s^2 2s^2 2p^6 3s^2 3p^5$

iv $1s^2 2s^2 2p^6 3s^2 3p^6$

Exercise 3.2 Ionisation energies and the Periodic Table

This exercise will familiarise you with the factors which influence the values of the first ionisation energy, it will also provide practice in interpreting graphs.

The graph shows the first ionisation energies, ΔH_{i1}, of successive elements A to M in the Periodic Table.

- Remember the three factors that influence 1st ionisation energy are: nuclear charge, distance of electrons from the nucleus and shielding.
- Think of these in terms of attraction between the negative outer electrons and the positive nucleus.
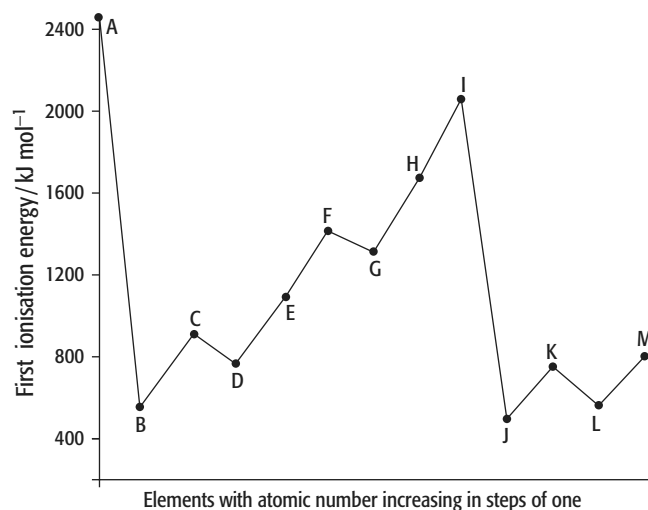
a Copy and complete these sentences using words from the list below.

attractive	**charge**	**electrons**	**increase**	**inner**
ionisation	**nucleus**	**principal**		

Across a period, there is a general _____ in the value of ΔH_{i1}. This is because of the increase in nuclear _____. Across a period the electrons are added to the same _____ quantum shell so the _____ forces between the _____ and the outer electrons increases gradually. So the first _____ energy increases gradually. Across a period there is not much difference in **shielding** because there are the same number of _____ shell _____.

b Which two elements A to M are noble gases?

c Between which two consecutive elements do these statements apply?

i When a new period starts, there is a sharp decrease in ΔH_{i1}. This is because the next electron added is in a **principal quantum shell** further from the nucleus.

ii When the atomic number increases by one, there is a decrease in ΔH_{i1} due to spin-pair repulsion.

iii When the atomic number increases by one there is a decrease in ΔH_{i1} because the next electron goes into a p subshell.

d Which two elements are in Group 2?

e Which element is in the same group as element M?

f Which element is in Group 17?

g What evidence is there from the graph that the value of ΔH_{i1} decreases down a group?

Exercise 3.3 Successive ionisation energies

This exercise will familiarise you with concept of successive ionisation energies, ΔH_{i1}, ΔH_{i2}, etc. and electron configurations. It will give you further practice in interpreting graphs.

The graph shows the successive ionisation energies of sodium plotted against the number of electrons removed.
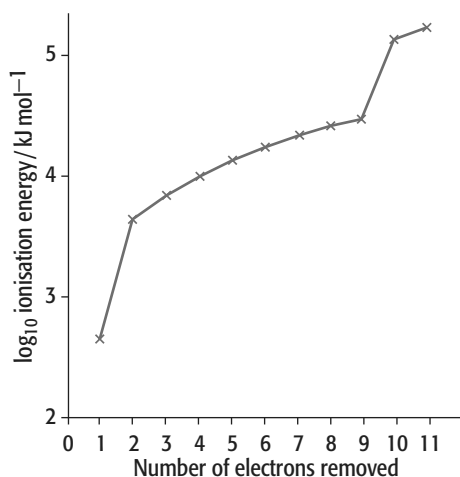
a How does this graph show that a sodium atom has:

    i one electron which is easily removed,

    ii two electrons in the first principal quantum shell,

    iii eight electrons in the second principal quantum shell?

When writing equations representing ionisation energies:

- The consecutive ionisation energies are represented by the charge on the ion formed, e.g. $Na^{4+}(g) \rightarrow Na^{5+}(g) + e^-$ represents the 5th ionisation energy.
- The ions formed are always positive and gaseous.

b The equation below represents the first ionisation energy of sodium.

$$Na(g) \rightarrow Na^+(g) + e^-$$

Write equations for:

    i The 4th ionisation energy of calcium.

   **ii** The 2nd ionisation energy of phosphorus.

Exercise 3.4 Electrons in orbitals

This exercise gives you further practice in deducing electron configurations and interpreting diagrams showing the direction of electron spin.

Remember when adding electrons to subshells that:

- Electrons are put one by one into separate orbitals in a given subshell spinning in the same direction.
- Electrons are only paired when no more empty orbitals in the subshell are available.

The diagram shows the arrangement of electrons in atomic orbitals and the direction of their spin.

2p [↑] [↑] [ ]

2s [↑↓]

1s [↑↓]
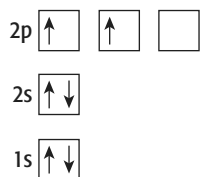
a How does this diagram show that:

    i in the second quantum shell, the electrons in p orbitals have more energy than the electrons in an s orbital,

    ii the electrons in an orbital spin in an opposite direction?

b Draw similar diagrams to show the arrangement and spin of electrons in atoms of:

    i oxygen

   **ii** chlorine

    iii phosphorus

Exercise 3.5 Interpreting line emission spectra

This exercise introduces you to how line emission spectra are used to obtain information about energy levels in atoms. It also provides practice in interpreting information.

This question contains some unfamiliar material. Think of the spectral lines being related to the energy levels.

Electrons are arranged in energy levels a certain distance from the nucleus of an atom. An electron can jump from a lower to a higher energy level when given a specific amount of energy (quantum of energy). When the electrons fall from higher to lower energy levels, radiation of particular frequencies is given off. The radiation can be recorded as lines in a line emission spectrum. Each line represents a particular frequency of radiation. The higher the frequency, the greater the energy associated with the radiation.

Part of a line emission spectrum is shown below, together with the electron energy levels.
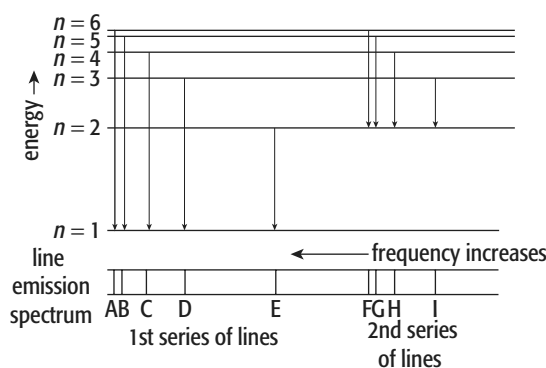
a   In the diagram, what do the letters *n* represent?

b   Describe how the spacing of each series of lines changes as the frequency increases.

c   What do the vertical arrows represent?

d   Which line in the spectrum represents the greatest energy change?

e   Which line in the spectrum represents the smallest energy change?

f   What happens to the distance between the energy levels as they get further from the nucleus?

g   Which line represents an electron falling from the 4th energy level to the 2nd energy level?

h   Use the information in the diagram to suggest why the first ionisation energy of lithium is much lower than the first ionisation energy of hydrogen.

Exam-style questions

QUESTION 1

Boron (atomic number 5) and gallium (atomic number 31) are in the same group in the Periodic Table.

a Write the electronic configuration using $1s^2$ notation for:

   i a boron atom [1]

   ii a gallium atom [1]

   iii a gallium ion. [1]

b Is a gallium ion larger than a gallium atom? Explain your answer. [1]

c Which atom, boron or gallium, has the lower first ionisation energy? Explain your answer. [3]

d **i** Explain why boron has a lower first ionisation energy than either beryllium or carbon. [6]

   ii Write an equation to represent the first ionisation energy of boron. [2]

e Gallium has both s and p-type orbitals. Draw the shape of each of these orbitals. [2]

Total: 17 marks

QUESTION 2

Chromium(III) chloride contains Cr^{3+} and Cl^- ions.

a Write the electronic configuration using $1s^2$ notation for:

   i a chloride ion [1]

   ii a chromium atom [1]

   iii a chromium(III) ion. [1]

b Name the element whose atoms have the same electronic configuration as a chloride ion. [1]

c Fluorine is in the same group as chlorine.

The graph below shows the $\log_{10}$ ionisation energy against the number of electrons removed from the fluorine atom.
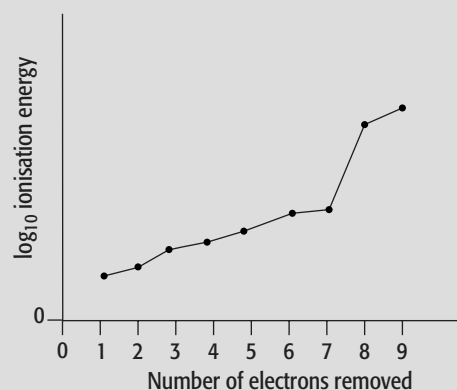
   i Explain why there is a gradual rise in the successive ionisation energies during the removal of the first 7 electrons. [3]

   ii Explain why there is a sharp rise in ionisation energy when the 8th electron is removed. [3]

   iii Write the equation which represents the 3rd ionisation energy of fluorine. [2]

Total: 12 marks

QUESTION 3

The table shows the first five successive ionisation energies of four metallic elements, A to D.

element	ΔH_{i1} / kJ mol^{-1}	ΔH_{i2} / kJ mol^{-1}	ΔH_{i3} / kJ mol^{-1}	ΔH_{i4} / kJ mol^{-1}	ΔH_{i5} / kJ mol^{-1}
A	578	1817	2745	11 578	14 831
B	496	4563	6913	9544	13 352
C	419	3051	4412	5877	7975
D	590	1145	4912	6474	8144

a Which two elements are in the same group of the Periodic Table? Explain your answer. [2]

b Which element requires most energy to convert one mole of its atoms to one mole of ions with a charge of 2+? [1]

c Which element, when it reacts, forms an ion with a charge of 3+? Explain your answer. [4]

d Suggest a value for the 6th ionisation energy of element C. Give a reason for your answer. [3]

e Vanadium is a transition metal.
   i Write the electronic configuration of vanadium using 1s^2 notation. [1]
   ii Write the equation which represents the second ionisation energy of vanadium. [2]

f An atom has the atomic number 16. Draw a diagram of electrons in their orbitals at different energy levels to show the electron arrangement in this atom. Show the spins of the electrons. [3]

Total: 16 marks

Chapter 4
Chemical bonding

Chapter outline

- Ionic, covalent and co-ordinate bonding using dot-and-cross diagrams
- Shapes and bond angles in molecules using ideas of electron-pair repulsion
- Covalent bonding in terms of overlap of atomic orbitals (σ-bonds, π-bonds and sp, sp^2 and sp^3 hybrid orbitals)
- Bond energy, bond length and bond polarity: using these to compare the reactivity of covalent bonds
- Intermolecular forces (permanent and induced dipoles and hydrogen bonding)
- Metallic bonding
- The effect of different types of bonding on physical properties
- Deducing the type of bonding from given information

KEY TERMS

Co-ordinate bond: Covalent bond in which both electrons come from the same atom.

Covalent bond: A pair of electrons shared between two atoms.

Dative covalent bond: See co-ordinate bond.

Delocalised electrons: Electrons not associated with a particular atom. They can move between three or more adjacent atoms.

Dot-and-cross diagram: The arrangement of outer shell electrons in a compound, the dots and crosses showing the origin of the electrons.

Electronegativity: The ability of an atom to attract the electrons in a covalent bond.

Hybridisation: The formation of a covalent bond by mixing different types of atomic orbitals.

Hydrogen bonding: Intermolecular forces between molecules having the structures RX—H and X—R′ where X is fluorine, nitrogen or oxygen and R or R′ is the rest of the molecule.

Intermolecular forces: Weak forces between molecules (van der Waals' forces).

Ionic bond: The electrostatic attraction between oppositely charged ions.

Lone pairs: Pairs of electrons in the outer shell of an atom not involved in bonding.

Non-polar (molecule): A molecule where the centres of positive and negative charge coincide.

Permanent dipole-dipole forces: Intermolecular forces between polar molecules.

Pi-bond (π-bond): A covalent bond formed by 'sideways' overlap of p and p or p and d atomic orbitals.

Polar (molecule): A molecule where the centres of positive and negative charge do not coincide, resulting in a dipole ($\delta^+ \longleftrightarrow \delta^-$) being formed.

Sigma bond (σ-bond): A single covalent bond formed by the 'end-on' overlap of atomic orbitals.

van der Waals' forces: Weak forces of attraction between molecules involving either temporary or permanent dipole-dipole forces.

Exercise 4.1 Ionic bonding and metallic bonding

This exercise will help you develop your understanding of ionic and metallic bonding.

a i Copy and complete the **dot-and-cross diagram** for magnesium fluoride. Show electrons originating from the metal atom as a cross (✗) and those originating from the non-metal as a dot (•).
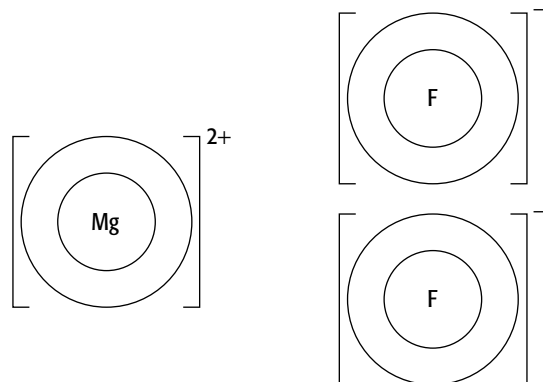
ii Draw dot-and-cross diagrams for

- lithium chloride
- sodium oxide.

In each case, show all the electrons and their shells.

Remember that in metallic structures:

- The particles present are metal ions and delocalised electrons.
- There are forces of attraction between the metal ions and the delocalised electrons.

b Copy and complete these sentences about bonding using words from the list below.

atoms   attraction   charge   covalent   directions   ions

An **ionic bond** is the strong force of electrostatic _____ between positive and negative _____ in a crystal lattice. The _____ on the ions is spread out in all _____. It is not like a _____ bond where the bonding is in one direction between two particular _____.

c Copy and complete these sentences which relate the properties of metals to their bonding.

move   ions   electrostatic   delocalised   electrons

Metals conduct electricity because the **delocalised electrons** are able to _____ throughout the structure between the positive metal _____. Many metals are strong and hard because the metal ions are held together by the strong _____ forces of attraction between the ions and the _____ _____.

## Exercise 4.2   Covalent bonding

This exercise will help you develop your understanding of covalent structures and how to draw dot-and-cross diagrams.

- A **covalent bond** is formed when a pair of electrons is shared between two atoms.
- When drawing more complex molecules such as ethanol, start with the carbon and oxygen atoms and try to join them together first by linking a single electron from one atom to a single electron from another.
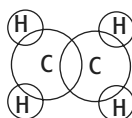
a Copy and complete these dot-and-cross diagrams for three covalent compounds. Show only the outer shell electrons.

lone pair
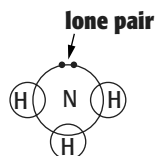
ammonia
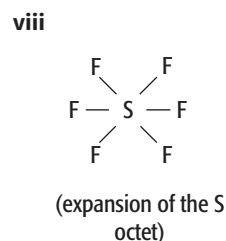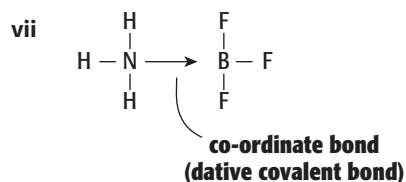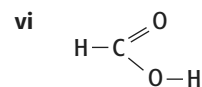show a H electron as x
show a N electron as •
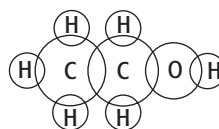
ethanol
show a H electron as x
show a C electron as •
show an O electron as ○

ethene
show a H electron as x
show a C electron as •

b Draw dot-and-cross diagrams for the molecules shown below. Show only the outer shell electrons. Note that the diagrams do not all show the exact shapes of the molecules.

i

P
Cl   Cl   Cl

ii

$N \equiv N$

iii

$H - Br$

iv

$Cl - S - Cl$

v

$O = C = O$

vi

$$H - C \overset{O}{\underset{O - H}{\diagdown}}$$

vii

$$H - N \longrightarrow B - F$$
H   H   F   F

**co-ordinate bond
(dative covalent bond)**

viii

F   F
F — S — F
F   F

(expansion of the S
octet)
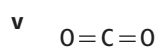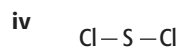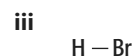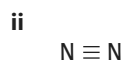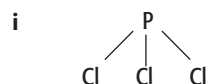
Exercise 4.3 Shapes of molecules

This exercise will help you deduce the shapes of molecules using the electron-pair repulsion theory.

Use this information to help you with the questions that follow:

stronger repulsion

lone pair of electrons repelling lone pair of electrons

lone pair of electrons repelling bonding pair of electrons

bonding pair of electrons repelling bonding pair of electrons

a Deduce the bond angles R to V in the following diagrams.

i

$$\overset{\cdots}{N}$$
H—R—H
H

ii

H
|
C
H—S—H
|
H

iii

Cl   Cl
B_T
|
Cl

iv

H
|
H—U—C
| V—Ö:
H |
H

b The diagrams below show the arrangement of the bonds in various molecules and ions. Draw and describe the shapes of these molecules or ions. On your diagrams give the values of the bond angle. Note that the diagrams do not show the exact shapes of the molecules and do not show lone pairs of electrons.
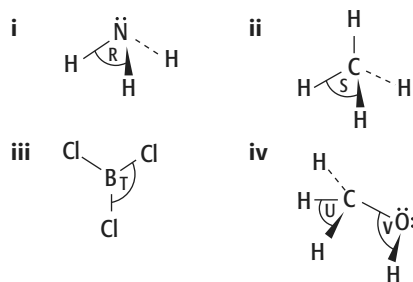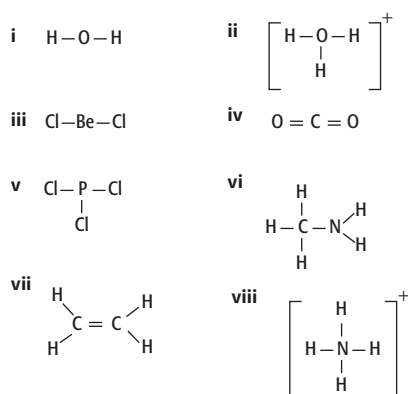
i   H—O—H

ii
$$\left[\begin{array}{c} H-O-H \\ | \\ H \end{array} \right]^{+}$$

iii Cl—Be—Cl

iv   O = C = O

v   Cl—P—Cl
|
Cl

vi
H
|
H—C—N
| H
H   H

vii

H   H
C = C
H   H

viii
$$\left[\begin{array}{c} H \\ | \\ H-N-H \\ | \\ H \end{array} \right]^{+}$$

Exercise 4.4 Intermolecular forces

This exercise will help you distinguish the relative strength of these forces based on molecular structure. It will also help you understand the relationship between these forces and the physical properties of simple molecules.

Remember that:

- **Electronegativity** increases from Group 1 to 17.
- Electronegativity increases up a Group.

a Put these atoms in order of their electronegativity. Put the most electronegative first.

chlorine   hydrogen   fluorine   nitrogen   oxygen

b i The diagram shows the electron clouds of hydrogen and hydrogen fluoride.

H   H
hydrogen

H   F
hydrogen fluoride
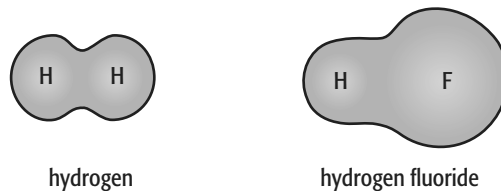
Copy these diagrams and put a + to show the centre of positive charge in each molecule and a − to show the centre of negative charge.

ii Explain why hydrogen fluoride is a **polar molecule**.

For similar sized molecules, the relative strength of the attractive forces between them are:

- **temporary dipole-induced dipole forces** (weakest)
- **permanent dipole-dipole forces**
- **hydrogen bonding** (strongest).

c Name the strongest type of attractive force between each of these pairs of molecules. Use the list above to help you.

  i CH_3Cl and CH_3Br

  ii CH_3NH_2 and CH_3OH

 **iii** $CH_3CH_2CH_2CH_2CH_2CH_3$ and $CH_3CH_2CH_2CH_2CH_3$

  iv CH_3COCH_3 and $CH_3COCH_2CH_3$

  v CH_3Br and CH_3NH_2

d Draw diagrams of each of these molecules to show how the atoms are arranged. On each diagram show the direction of the dipole as $\delta^+ \longleftrightarrow \delta^-$. If no net dipole is present, write 'none'.

  i CH_2Cl_2

  ii CBr_4

  iii NH_3

  iv $ClBr$

e Suggest in terms of intermolecular forces why:

  i water has a higher enthalpy change of vaporisation than pentane, C_5H_{12}, even though pentane has a higher molar mass.

  ii pentane is a liquid at r.t.p. but butane C_4H_{10} is a gas at r.t.p.

  iii CH_3NH_2 has a higher boiling point than CH_3Cl.

Exercise 4.5 Bonding and orbitals

This exercise focuses on covalent bonding, the hybridisation of atomic orbitals and how these affect the properties of simple molecules.

Remember that:

- Hybridisation of orbitals is a mixing of s and p orbitals depending on the types of orbitals available.
- Graphite has layers of hexagonally arranged carbon atoms.

a Atomic orbitals overlap in various ways. Three types of hybrid orbital are sp, sp^2 and sp^3.

  i The diagrams below show the hybridised orbitals in ethane and ethene. What words best represent the letters A to D?

A bond

Overlap of C hybrid orbitals

B bond   A bond

Overlap of three D hybrid orbitals
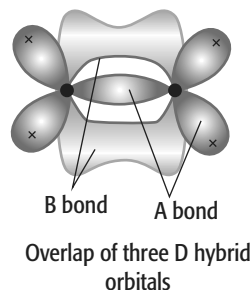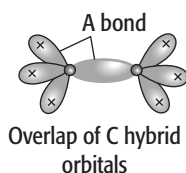
  ii Refer to the hybridised structure of ethane to explain why the H−C−H bond angles are 109.5º.

b The diagram below shows the p orbitals in benzene, C_6H_6.
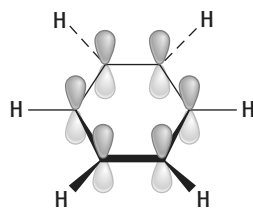
i Predict the shape of the electron clouds arising from the *p* electrons when the *p* electrons are brought closer to each other.

ii Suggest why some of the electrons in benzene are delocalised.

iii Benzene does not conduct electricity even though it has delocalised electrons. Explain why it does not conduct.

iv Use ideas of overlap of p orbitals to suggest why graphite conducts electricity.

c When molecules react, bonds are broken and new bonds are formed. Suggest why oxygen is more reactive than nitrogen.

d The type of bond, **sigma** or **pi**, influences chemical reactivity. Suggest why ethene is more reactive than ethane.

e The polarity of a bond can influence chemical reactivity. Suggest why $CH_3CH_2CH_2Cl$ is more reactive than $CH_3CH_2CH_3$.

Exam-style questions

QUESTION 1

This question is about the structure of simple molecules and the relative strength of the forces between these molecules.

- Make sure that you can identify the atoms responsible for hydrogen bonding and permanent dipole-dipole forces.
- The command word *describe* in part (a) means that you have to write down the pattern that you see.
- The command word *explain* means that you have to give scientific reasons for your answer.

The table shows the boiling points of some hydrides.

hydride	CH_4	SiH_4	GeH_4	SnH_4	NH_3	H_2O	HF
boiling point / K	112	161	185	221	241	373	293

a Describe and explain the trend in the boiling points of the Group 14 hydrides CH_4, SiH_4, GeH_4 and SnH_4 [4]

b Explain the differences in the period 2 hydrides CH_4, NH_3, H_2O and HF in terms of attractive forces between the molecules. [6]

c Draw a dot-and-cross diagram for water. [2]

d i Water is a V-shaped molecule. Explain why. [2]

ii Suggest a value for the H—O—H bond angle in water. [1]

e CH_3Cl is a polar molecule. Draw a diagram to show the three-dimensional structure of this molecule. On your diagram show the direction of the dipole as $\delta^+ \longrightarrow \delta^-$. [2]

f Suggest why CCl_4 is not a polar molecule. [1]

Total: 18 marks

QUESTION 2

Polymers consist of long chains of molecules. When heated, the chains begin to move over each other. The temperature at which this happens is called the glass point, T_g.

The structures of a section of the chains of three polymers are shown below together with their T_g values.

Polymer A $-CH-CH_2-CH-CH_2-$ $T_g = 378\,K$
                      | |
                      CN   CN

Polymer B $-CH-CH_2-CH-CH_2-$ $T_g = 353\,K$
                      | |
                      Cl   Cl

Polymer C $-CH-CH_2-CH-CH_2-$ $T_g = 298\,K$
                      | |
                      CH_3 CH_3
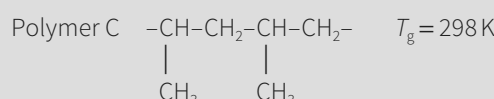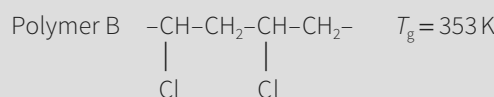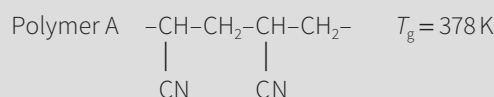
a Explain these differences in the values of T_g by referring to the relative strengths of permanent and temporary dipole-dipole forces. [5]

b The CN group is called a nitrile group. This group has a triple bond. Draw a dot-and-cross diagram of the nitrile ion CN^-. Show only the outer shell electrons. [2]

c Liquid ammonia contains molecules, NH_3, which are extensively hydrogen bonded. Draw a diagram to show two hydrogen-bonded molecules of ammonia. Show the lone pairs of electrons and the correct orientation of the hydrogen bond. [3]

d i Aqueous ammonia reacts with dilute sulfuric acid to form a salt. Write the balanced equation for this reaction including state symbols. [3]

ii The salt formed in this reaction boils (under pressure) at 513 °C. Ammonia boils at −33 °C. Explain the difference between these boiling points in terms of bonding. [4]

e Magnesium oxide is also a salt. Draw a dot-and-cross diagram of the ions in magnesium oxide. Draw only the outer shell electrons. [2]

Total: 19 marks

QUESTION 3

a The structure of iodic acid is shown below.

    O⫽ I ⫽O
        |
        O
        |
        H
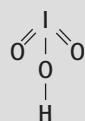
i Draw a dot-and-cross diagram for iodic acid. [3]
ii Explain why iodic acid has this pyramidal structure. [2]
iii Suggest a value for the O—I—O bond angle. [1]

b Iodine is a solid at r.t.p. but hydrogen iodide is a gas at r.t.p. Explain this difference in terms of intermolecular bonding. [2]

c Hydrogen iodide has a much lower boiling point than hydrogen fluoride. Explain why in terms of intermolecular bonding. [3]

d The structural formulae of 1-iodopropane and 2-iodopropane are shown below.

CH_3—CH_2—CH_2—I CH_3—CHI—CH_3
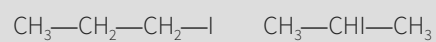
1-iodopropane   2-iodopropane

Suggest why 2-iodopropane has a lower boiling point than 1-iodopropane. [3]

e At low temperatures, aluminium chloride, $AlCl_3$, forms a molecule with the structure shown below.

Cl   Cl   Cl
\ / \ /
Al   Al
/ \ / \
Cl   Cl   Cl

Name the type of bond shown by the arrows and suggest why $AlCl_3$ molecules join together in this way. [2]

Total: 16 marks
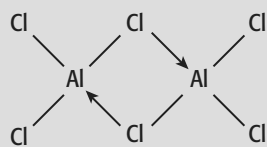
Chapter 5
States of matter

Chapter outline

- Solid state in terms of motion, arrangement and proximity (distance apart) of particles
- The lattice structures of crystalline solids (simple molecular, giant molecular, giant ionic and metallic)
- Liquid state in terms of the motion, arrangement and proximity of particles
- Melting, vaporisation and vapour pressure
- Suggesting from given physical data, the structure and bonding present in a substance
- Assumptions of the kinetic theory of ideal gases
- Ideal gases in terms of intermolecular forces and molecular size
- Using the general gas equation in calculations, including the determination of relative molecular masses
- The finite nature of materials as a resource and the importance of recycling

KEY TERMS

Allotrope: Different crystalline or molecular forms of the same element, e.g. graphite and diamond.

Buckminsterfullerene, C_{60}, or hollow tubes **(nanotubes).**

Fullerenes: Molecular allotropes of carbon in the form of hollow spheres, e.g. buckminsterfullerene.

Giant ionic structure: A structure having a lattice of positive and negative ions.

Giant molecular structure: A structure having a three-dimensional network of covalently bonded atoms throughout the whole structure. There are no separate molecules.

Graphene: A single sheet of hexagonally arranged carbon atoms.

Ideal gas: A gas whose volume increases proportionally with increase in the temperature (in Kelvin) and whose volume is inversely proportional to pressure.

Ideal gas equation: $pV = nRT$ (where p is pressure, V is volume, n is the number of moles of gas, R is the gas constant and T is the temperature (in Kelvin).

Kinetic theory: The theory that particles in gases and liquids are in constant motion, and in solids the particles are vibrating around a fixed point.

Lattice: A regularly repeating arrangement of atoms, molecules or ions in three dimensions throughout the whole structure.

Metallic structure: A structure having a lattice of metal ions surrounded by delocalised electrons.

Simple molecular structure: A structure containing identical molecules which are separated from each other by weak forces.

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

Vapour pressure: The pressure exerted by a vapour when it is in equilibrium with the liquid from which it vaporised.

Exercise 5.1 Properties of the three states of matter

This exercise will familiarise you with the properties of the three **states of matter** and the structure of a simple covalent molecule and a **giant molecular** (giant covalent) **structure.** It will also help you relate the properties of these substances to their structures.

- Structures can be described as simple molecular or giant and bonding can be covalent, ionic or **metallic**.
- You can combine these for a complete description of a compound or element, e.g. simple covalent molecule, **giant ionic structure**, giant covalent structure.

a Carbon dioxide, bromine and sulfur have **simple molecular structures**. Copy and complete the table to show the proximity (closeness), arrangement and motion of the molecules in these three substances. Three of the answers have been done for you.

substance	state at 25°C and 1 atm pressure	proximity	arrangement	motion
bromine	liquid		irregular / random	
carbon dioxide	gas			moving rapidly from place to place
sulfur	solid	close together		

b The structure of solid carbon dioxide and part of the structure of one form of boron nitride are shown below.

Carbon dioxide
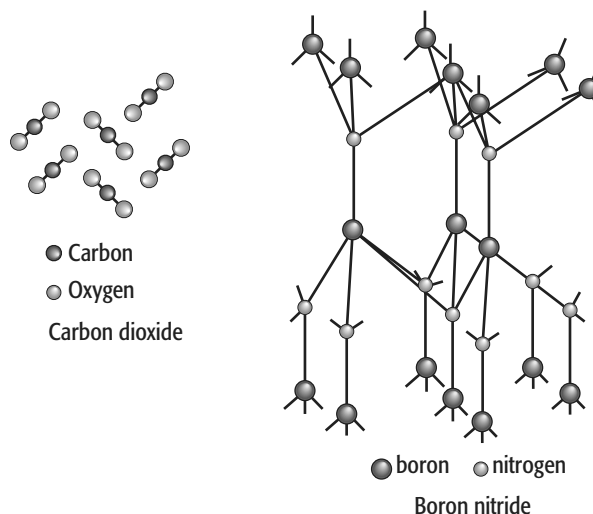
boron   nitrogen

Boron nitride

Figure 5.01 The structures of carbon dioxide and boron nitride

i Copy the following passage selecting the correct statements in bold about the structure and properties of boron nitride.

Boron nitride has a **giant / simple** molecular structure. It has **covalent / ionic** bonds. Its structure is similar to the structure of **diamond / graphite**. Each boron atom is bonded to **three / four** nitrogen atoms. The nitrogen atoms are arranged **hexagonally / tetrahedrally** around each boron atom. Boron nitride has a **high / low** melting point because all the **atoms / molecules** are joined by **strong / weak** bonds. Boron nitride **conducts / does not conduct** electricity because all the outer electrons in the boron and nitrogen atoms are used in bonding.

ii Copy and complete these statements by writing suitable words or phrases in the spaces.

Carbon dioxide has a _____ _____ structure. It has _____ _____ bonding between each carbon and oxygen atom. Carbon dioxide is a gas at room temperature because the forces _____ .

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Exercise 5.2 Giant structures

This exercise will help you distinguish between the three types of giant structure: ionic, giant molecules (giant covalent structures) and metals. It will also help you relate the properties of these substances to their structures.

Diagrams can give you important information about structure and bonding.

- + and − signs in circles indicate ions but if a − sign is not in a circle it will indicate an electron.
- Giant covalent structures usually have continuation bonds shown to indicate that the structure is only part of a larger structure.

The diagram shows three different types of giant structure, **A**, **B** and **C**.

| | A | B | C |
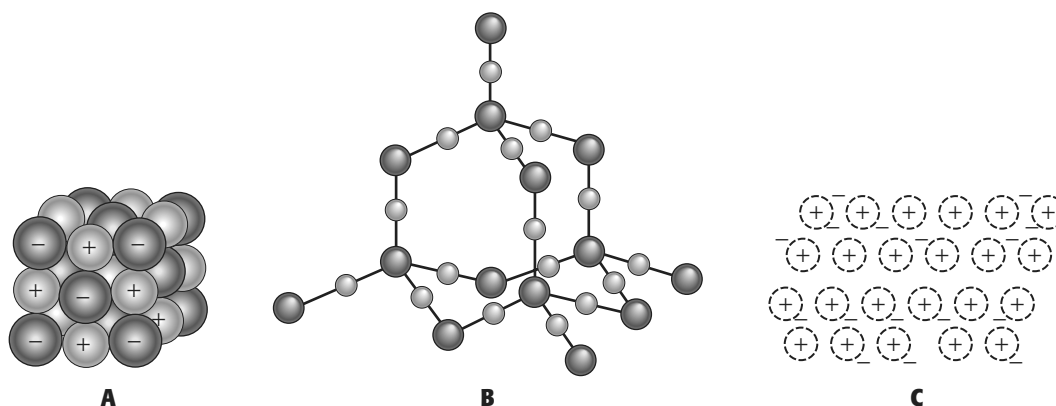
a Name each of the three types of giant structure represented by the letters **A**, **B** and **C**.

b Copy and complete the table to compare the structures and properties of **A**, **B** and **C**.

	structure **A**	structure **B**	structure **C**
Type of particles present in the diagram		atoms of Si and O	
Melting point	high		(generally) high
Electrical conductivity of solid		does not conduct	
Electrical conductivity when molten	conducts		

c Match the properties 1 to 5 of compounds A and B on the left to the correct explanations A to E on the right.

Property	**Explanation**
1. Compound A has crystals with a regular shape	A. because the ions are free to move from place to place
2. Compound B does not conduct electricity when molten	B. because there are strong attractive forces between the large number of positive and negative ions
3. Compound A is hard	C. because there are neither ions nor electrons free to move throughout the structure
4. Compound B does not dissolve in water	D. because the particles are arranged in a **lattice**
5. Compound A conducts electricity when molten	E. because the forces of attraction between the atoms of B are greater than the forces of attraction between the atoms of B and the water molecules

d Describe the structure of metals and explain why they conduct electricity and are malleable. In your answer include the following words and phrases:

outer shell electrons   **delocalised electrons**   **metal ions**

layers of metal ions   **attractive forces between the metal ions and delocalised electrons**

Exercise 5.3 Vapour pressure and molar mass of gases

This exercise gives you practice in revising the concept of **vapour pressure** and using the **ideal gas equation, $pV = nRT$** to calculate relative molecular mass. It also develops your skills in processing data from the results of an experiment.

- Vapour pressure is an example of physical equilibrium. You may want to revise the concept of equilibrium (Chapter 8) before doing part (a).
- When using the equation $pV = nRT$ make sure that the units are the correct ones, e.g. temperature in K, volume in m^3.

a i Copy and compete the following sentences about vaporisation using words from the list below. Some words may be used more than once.

condensing   equilibrium   evaporates   liquid   pressure   vapour   weaker

When a volatile liquid is heated or left to evaporate, the attractive forces between the molecules get _____. Some molecules escape to form a _____. A container is half-filled with liquid hexane, C_6H_{14}. Some of the hexane _____ until an _____ is reached between the hexane

molecules escaping from the _____ and molecules of hexane vapour _____ back to liquid hexane. The pressure exerted by the hexane vapour in equilibrium with its liquid is called its _____ _____.

ii Write an equation showing the equilibrium between liquid hexane and hexane vapour.

b The relative molecular mass of a volatile liquid, **L**, can be found experimentally using the apparatus shown below.
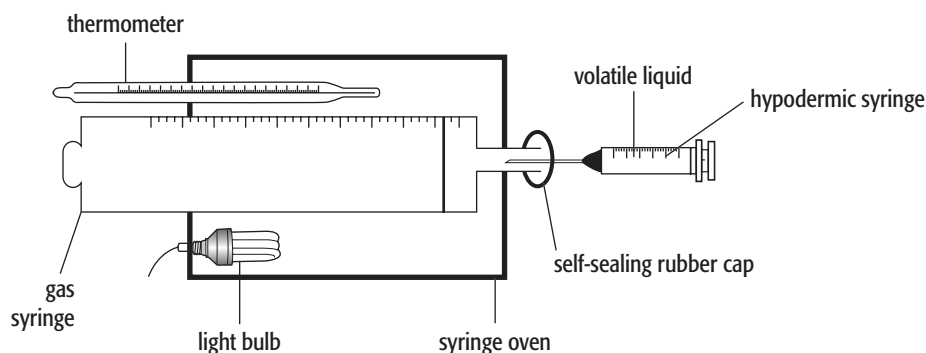
The results are given below:

Temperature in the syringe oven = 120 °C

Mass of hypodermic syringe at the start of the experiment = 10.71 g

Mass of hypodermic syringe after injection of **L** into gas syringe = 10.54 g

Volume of air in the gas syringe at the start of the experiment (at 120 °C) = 4.0 cm³

Volume of vapour in the gas syringe after vaporisation of the liquid = 69.0 cm³

Atmospheric pressure = 1.00×10^5 Pa

i Use the information to determine the following values:
   The temperature of the vapour = _____ K
   The mass of **L** vaporised = _____ g
   The volume of vapour in the gas syringe in m³ = _____ m³

ii Rearrange the gas equation $pV = nRT$ to make n (number of moles of vapour) the subject.

iii Calculate the value of n ($R = 8.31$ J K⁻¹ mol⁻¹). Express your answer to 2 significant figures.

iv Use the value from part **iii** and the mass of **L** injected into the gas syringe to calculate the relative molecular mass of **L**. Express your answer to 2 significant figures.

v What effect, if any, would the following errors make on the measured value of the relative molecular mass? In each case, explain your answer.
   • Having 4 cm³ of air (at 120 °C) in the syringe at the start of the experiment.
   • Losing some of liquid **L** from the syringe into the atmosphere during its injection.
   • Reading the temperature inside the syringe oven as 130 °C instead of the correct 120 °C.

Exercise 5.4 Aluminium smelting and recycling

This case study about an aluminium smelting plant introduces you to some environmental aspects of metal extraction and the importance of recycling.

You should be prepared to answer questions relating to economic or environmental aspects of chemistry. In this example you are given enough information to enable you, together with your own knowledge, to give good answers.

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a Read the extract about the manufacture of aluminium and answer the questions that follow.

Aluminium is manufactured by the electrolysis of molten aluminium oxide. The aluminium oxide is extracted from bauxite ore.

Aluminium oxide is purified using concentrated sodium hydroxide then electrolysed at 1000 °C using a very high electric current (40 000 amps). The process produces a large amount of heat energy, so a lot of water is needed for cooling purposes. The cells ('pots') have to be renewed from time to time and the poisonous arsenic which builds up in the pots has to be disposed of carefully. Aluminium is also recycled from old car bodies, cooking pots and drinks cans. The aluminium is melted in a furnace and 'degassed' to remove hydrogen. Recycling aluminium uses a relatively small amount of energy compared with extracting aluminium using electrolysis but small amounts of poisonous materials may build up in the furnace.

The Government of Trinidad and Tobago in the Caribbean hoped to set up an aluminium smelting works close to the small coastal town of La Brea. La Brea has a poorly developed road system. A new harbour would be built to import bauxite from Jamaica or Guyana, which are also in the Caribbean. There is other industry in the area. Aluminium would also be recycled in the area. The aluminium smelter was not built due to a variety of reasons.

i Suggest three reasons why the government of Trinidad and Tobago might want to build an aluminium smelting works at La Brea.

ii Suggest three reasons why the aluminium smelter was not built.

iii Give three reasons why recycling aluminium is better for the environment than extracting aluminium from bauxite.

b The graph shows the percentage of aluminium cans and plastic bottles recycled over a 20-year period.
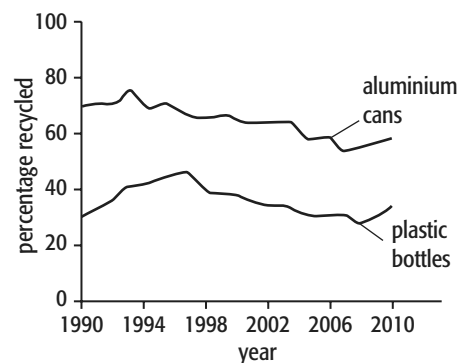

Figure 5.02 The percentage of aluminium cans and plastic bottles recycled from 1990 to 2010

What are the similarities and differences in the percentage of aluminium cans and plastic bottles recycled from 1990 to 2010?

Exercise 5.5 Allotropes of carbon

This exercise will help you investigate some **allotropes** of carbon including **fullerenes** and **graphene**.

Diamond, graphite, fullerenes and graphene are all allotropes of carbon.