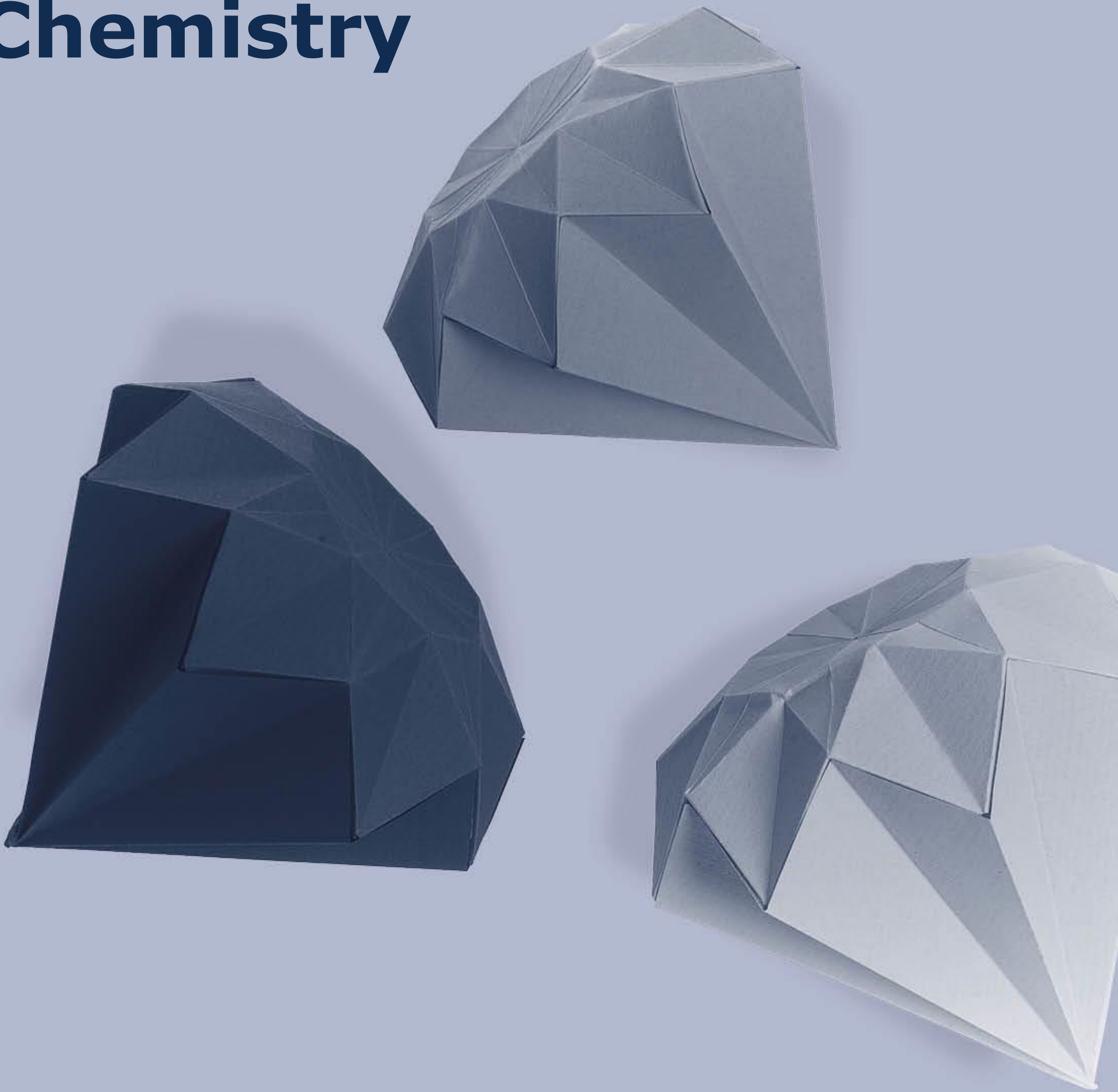
edexcel

AS and A Level Chemistry



TRANSITION GUIDE

Reinforcing knowledge, skills and literacy in chemistry

ALWAYS LEARNING

PEARSON

Contents

Introduction

Reinforcing knowledge, skills and literacy in chemistry

From our research, we know that it is easy for teachers to fall into the trap of going over work that has already been covered extensively at KS4. This may be because of a feeling that during the summer break students have forgotten what they had been taught or, if they are from different centres, uncertainty about the standard they have reached so far. This is where you can lose valuable teaching time and later find yourself rushed to complete the A-level content.

To help you with planning and teaching your first few A-level lessons and to save you time, we have worked with practising teachers and examiners to develop these valuable, focused transition materials. These will help you reinforce key concepts from KS4 and KS5 and guide your students’ progression.

These transition materials include:

- mapping of KS4 Edexcel GCSE(s) to the new Edexcel A Level Chemistry specifications
- baseline assessments
- summary sheets
- student worksheets
- practice questions.

The mapping of content and skills from KS4 to KS5 should enable you to streamline your teaching and move on to the KS5 content within the first two weeks of term.

This will serve two purposes.

1 Learners will feel they are learning something new and will not get bored with over-repetition – particularly true for your most able learners.

2 Learners will be able to discover very early on in the course whether A level chemistry is really a suitable subject choice for them.

You may choose to use this resource in one of several ways.

- After KS4 exams – if your school brings back Yr11 learners after their exams.
- In sixth-form induction weeks.
- As summer homework in preparation for sixth form.
- To establish the level of performance of your students from their range of KS4 qualifications.

Transition guide overview

Topic	Specification links	Resources
Section A Atomic structure, formulae and bonding	KS5 **–** Topic 1 **–** Atomic structure and the Periodic Table KS4 – Core and Additional concepts	• Students' strengths and misconceptions • Building knowledge • Summary sheets • Worksheet 1: Atomic structure and the Periodic Table • Worksheet 2: Orbitals and electron configuration • Exam report and discussion • Exam practice
Section B Quantitative analysis and equations	KS5 **–** Topic 5 **–** Formulae, equations and amounts of substance KS4 **–** Additional and Further Additional/Extension concepts	• Students' strengths and misconceptions • Building knowledge • Summary sheet: Writing formulae • Worked examples: Calculations • Worksheet 1: Chemical formulae • Worksheet 2: Cations and anions • Worksheet 3: Writing equations • Exam practice

Topic	Specification links	Resources
Section C Structure and properties **–** Literacy Focus	KS5 **–** Topic 2 **–** Bonding and Structure KS4 **–** Core and Additional concepts	• Students' strengths and misconceptions • Building knowledge • Summary sheet 1: Ionic structure and bonding • Summary sheet 2: Diamond and graphite structure • Teaching ideas: Using key words to describe ionic structure • Exam practice
Appendix 1	Specification mapping	
Appendix 2	Further baseline assessment questions	

The table below outlines the types of resources to be found in each section along with a description of its intended uses.

Type of resource	Description
Baseline assessment	This tests fundamental understanding of: • atomic structure • electron configuration (2.8…) • dot-and-cross diagrams for covalent and ionic compounds • definitions of types of bonding; distinguishing between bonding and structure; explaining properties in terms of bonding.
Students' strengths and misconceptions	Students' strengths and common misconceptions.
Building knowledge	May be used to assess understanding and for reflection on learning. Used for setting targets for improvement.
Summary sheets	Review of KS4 concepts. Summary of key points and guide to correct use of key terms. Tips on how to answer exam questions.
Student worksheets	Checking understanding of key points from Baseline assessment and Summary sheet. Checking understanding of new KS5 learning.
Exam practice and Examiners' report	How to answer exam-type questions and KS5 level.

Baseline assessment

Name: ______________________________ Form: ______________

Chemistry group: ______________________________

GCSE Chemistry/Science grade: ______________________

Date: ______________________________

Targets for improvement

- ☐ Writing formulae
- ☐ Naming compounds
- ☐ Atomic structure
- ☐ Electron configuration
- ☐ Word equations
- ☐ Balancing equations
- ☐ Definition of bonds

Question	**Marks**
1	/4
2	/5
3	/3
4	/4
5	/5
6	/15
7	/6
8	/6
9	/4
Total	/52
%	
Grade	

Target grade

- ☐ **OT**
- ☐ **BT**
- ☐ **AT**

1 Give the formulae of the following compounds.

Copper(II) sulfate ______________________

Lithium hydrogencarbonate ______________________

Sodium hydroxide ______________________

Potassium nitrate ______________________

Strontium nitrate ______________________

Calcium hydroxide ______________________

Sodium carbonate ______________________

Aluminium fluoride ______________________

(4 marks)

2 Name the following compounds.

NH_4Cl ______________________

HNO_3 ______________________

C_2H_4 ______________________

C_3H_8 ______________________

CO_2 ______________________

C_2H_5OH ______________________

Fe_2O_3 ______________________

SO_2 ______________________

HBr ______________________

NH_3 ______________________

(5 marks)

3 Complete the table below.

Particle	Where it is found	Charge	Mass
		0	
Proton			
			0

(3 marks)

4 Deduce the relative formula mass of the following.

SO_2	____________________	KBr	____________________
C_2H_6	____________________	$Ca(OH)_2$	____________________
C_2H_5OH	____________________	$NaNO_3$	____________________
NH_4Cl	____________________	$FeCl_3$	____________________

(4 marks)

5 State what is meant by the following terms.

a the mass number of an atom

(1 mark)

b relative atomic mass

(2 marks)

c isotopes

(2 marks)

6 For the following reactions, write:

a the word equation (1 mark)

b the chemical equation complete with state symbols. (2 marks)

Calcium carbonate and hydrochloric acid

Magnesium and sulfuric acid

Complete combustion of butane

Thermal decomposition of calcium carbonate

Sodium and water

(12 marks)

7 State what is meant by the following terms.

Ionic bonding

Covalent bonding

Metallic bonding

(3 marks)

8 Complete the table below. You may use the following words to help you.

ionic covalent giant simple metallic

Substance	**Formula**	**Type of bonding**	**Type of structure**
Hydrogen sulfide			
Graphite			
Silicon dioxide			
Methane			
Calcium			
Magnesium chloride			

(6 marks)

9 Explain why graphite can be used as a solid lubricant and also as electrodes.

(4 marks)

-End of assessment-

Section A: Atomic structure, formulae and bonding

This section reviews the fundamental concepts from Core and Additional Science. The resources provide a progressive journey, from simple knowledge of the subatomic particles to the more complex electron arrangements in orbitals. It is important to emphasise that the AS concepts are amplifications of what was learnt at KS4. There are opportunities for students to review KS4 work to strengthen their foundation and for teachers to bring their teaching groups together to the same starting level.

Students' strengths and common misconceptions

The table below outlines the areas in which most students do well and the common mistakes and misconceptions across the topics listed.

	Strengths	**Common mistakes**
Atomic structure	Listing subatomic particles and their properties (mass and charge).	Being unclear about Subatomic particles in ions.
Electron configuration	Simple 2.8.8… rule.	Not realising that the s, p, d configuration is an amplification of the 2.8.8… format. Deducing group number for the p-block elements (e.g. group7 – not counting the s-electrons with the p-electrons as outer electrons). Misunderstanding electron configuration for ions. Confusing the terms 'orbital' and 'energy level'.
Dot-and-cross diagrams	Knowing the general rule for individual atoms. Simple ionic compounds e.g. NaCl.	Checking the total outer electrons after bonding – both ionic and covalent. Overlapping shells for ionic compounds. Missing charges on ions.

Table of resources in this section

Topics covered	Type of resource	Resource name	Brief description and notes for resource
• Atomic structure and formulae • Electronic configuration	Teacher resource	Building knowledge	Building knowledge learning outcomes. May be used to assess understanding and for reflection on learning. Used for setting targets for improvement.
• Atomic structure • Ionic compounds • Electron configuration • Dot-and-cross diagrams for ionic bonding • Covalent compounds (simple covalent bonding)	Teacher resource	Summary sheets	Review of KS4 concepts. Summary of key points and guide to correct use of key terms. Tips on how to answer exam questions.
• Atomic structure and the Periodic Table	Student worksheet	Worksheet 1: Atomic structure and the Periodic Table	Checking understanding of key points from Baseline assessment and Summary sheet.
• Orbitals and electron configuration	Student worksheet	Worksheet 2: Orbitals and electron configuration	Checking understanding of new KS5 learning.
• Definition of isotopes • Atomic number and relative isotopic mass • Dot-and-cross diagrams for ionic and covalent bonds	Exam report and discussion	Examples of students’ responses from Results Plus – Examiners’ report	How to answer exam-type questions at KS5 level. Covering main misconceptions for main topics.
• Writing formulae • Atomic structure • Electron configuration • Dot-and-cross diagrams	Student questions	Exam practice	Exam questions on section covering KS4 to KS5 content. Checking how far students have progressed at the end of the section.

Building knowledge

Atomic structure and formulae

Building your understanding

Deduce formulae of compounds with compound cations and/or anions.

Use dot-and-cross diagrams to show ionic and covalent bonding.

Write chemical formulae with two elements.
Know the sign on ions made up of single elements.
Deduce the subatomic particles in ions.

Deduce the components of an atom from its atomic and mass number.
Recall the subatomic particles and know the mass and charge of each.

Electron configuration

Building your understanding

You will need to know:
- the order of filling
- how many electrons in each orbital.

If you are given the atomic number, you should be able to show how the electrons are arranged in their orbitals: s, p and d.

You should also be able to represent these using ‘electrons in boxes’.

From KS4, know the rules for electron configuration: demonstrate this in dot-and-cross diagrams and in the shorthand form e.g. 2.8.8

KS4 – Atomic structure

Subatomic particles: nucleus (protons and neutrons), electrons in shells.

Describe the particles in terms of their relative masses and relative charges:

- Protons – mass 1, charge +1.
- Electrons – mass = negligible ($\frac{1}{1840}$), charge −1.
- Neutrons – mass = 1, charge = 0.

Notes

- Number of protons = number of electrons (uncharged/neutral atoms).
- Proton number = atomic number.
- Mass number = protons + neutrons.

KS4 – Isotopes and calculating relative isotopic mass

Isotopes are *atoms* of the same elements which have different numbers of *neutrons* but the same number of *protons*.

$$\text{Relative isotopic mass} = \frac{\text{sum of (\% abundance} \times \text{ isotopic mass)}}{100}$$

KS4 – Ionic compounds

Formation of ions

Atoms of metallic elements in Groups 1,2 and 3 can form positive ions when they take part in reactions since they are readily able to lose electrons.

Atoms of Group 1 metals lose one electron and form ions with a 1+ charge, e.g. Na^{+}

Atoms of Group 2 metals lose two electrons and form ions with a 2+ charge, e.g. Mg^{2+}

Atoms of Group 3 metals lose three electrons and form ions with a 3+ charge, e.g. Al^{3+}

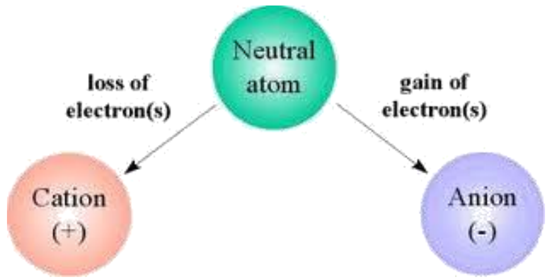
Atoms of non-metallic elements in Groups 5, 6 and 7 can form negative ions when they take part in reactions since they are able to gain electrons.

Atoms of Group 5 non-metals gain three electrons and form ions with a 3– charge, e.g. N^{3-}

Atoms of Group 6 non-metals gain two electrons and form ions with a 2– charge, e.g. O^{2-}

Atoms of Group 7 non-metals gain one electrons and form ions with a 1– charge, e.g. Cl^{-}

ANions = Negative Ca+ions = +ive



Why are ions negative or positive?

- Find the atomic number (the smaller number with the symbol).
- This equals the number of protons, which equals the number of electrons in an uncharged/neutral atom.
- If electrons are lost from the atom, there are now more protons than electrons, so the ion is positively charged.
- If electrons are gained by the atom, there are now fewer protons than electrons, so the ion is negatively charged.

KS4 – Electron configuration

Filling electron shells

- n = 1, maximum = $2e^-$
- n = 2; maximum = $8e^-$
- n = 3 ;maximum = $18e^-$
- n = 4; maximum = $32e^-$

Representing electron configurations

- Write as e.g. 2.8.3 or 2,8,3

Using the Periodic Table

- Period number (row) = number of shells
- Group number (column) = number of electrons in the outer (last) shell

Group number	**1**		**2**		**3**				**5**		**6**		**7**	
	Li		**Be**		**B**				**N**		**O**		**F**	
	Atom	Ion	Atom	Ion	Atom	Ion			Atom	Ion	Atom	Ion	Atom	Ion
Electrons	−3	−2	−4	−2	−5	−2			−7	−10	−8	−10	−9	−10
Protons	+3	+3	+4	+4	+5	+5			+7	+7	+8	+8	+9	+9
Overall charge	0	1+	0	2+	0	3+			0	3−	0	2−	0	1−
Electron configuration	2.1	2	2.2	2	2.3	2			2.5	2.8	2.6	2.8	2.7	2.8
Name of ions	lithium		beryllium		boron				nitride		oxide		fluoride	
	Lose electrons, charge = +group number								Gain electrons, charge = group number − 8					

KS4 – Dot-and-cross diagrams for ionic bonding

Hints and tips

Always …

… count the electrons!

… remember that ions should have full outer shells.

… make sure that when an ion is formed, you put square brackets round the diagram and show the charge.

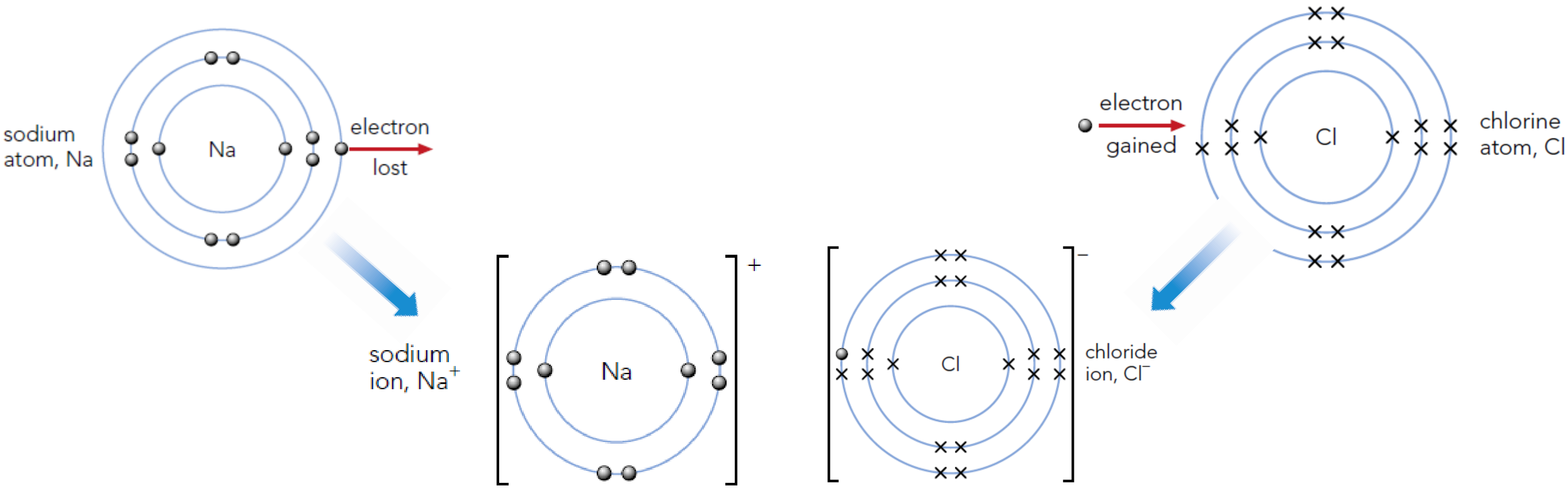
Never …

… show the electron shells overlapping.

… show electrons being shared (ions are formed by the **transfer** of electrons!).

… remove electrons from the inner shell.

… give metals a negative charge.



KS4 – Covalent compounds (simple covalent bonding)

A covalent bond is form when a pair of electrons is shared between two atoms.

Covalent bonding results in the formation of molecules.

Hints and tips

Always …

… show the shells touching or overlapping where the covalent bond is formed.

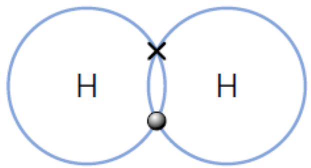
… count the final number of electrons around each atom to make sure that the outer shell is full.

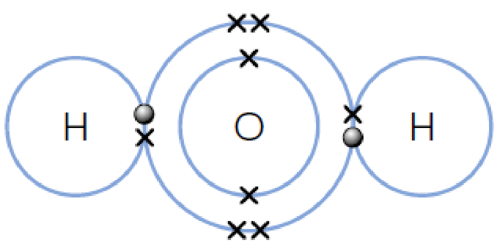
Never …

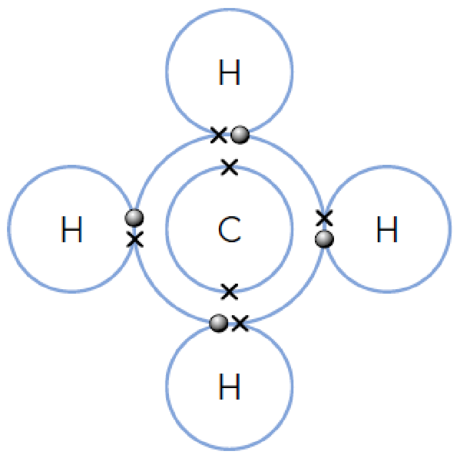
… include a charge on the atoms.

… draw the electron shells separated.

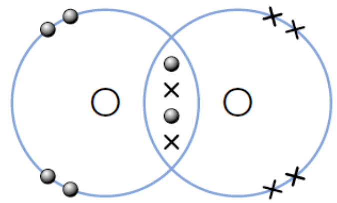
… draw unpaired electrons in the region of overlap.

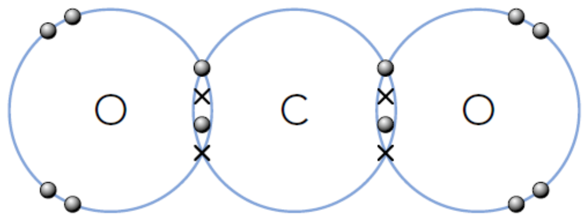






The two diagrams below only show the outer-shell electrons.





Worksheet 1: Atomic structure and the Periodic Table

Complete the following sentences and definitions to give a summary of this topic.

Structure of an atom

The nucleus contains …

The electrons are found in the …

To work out the number of each sub-atomic particle in an atom we use the Periodic Table (PT). The number of protons is given by …

In a neutral atom the number of electrons is …

To work out the number of neutrons we …

Vocabulary

State what is meant by the following terms.

1 Relative atomic mass

2 Relative molecular mass

3 Isotope

4 Relative isotopic mass

Structure of an ion

When an atom becomes an ion, only the number of ______________________ changes.

For positive ions this ________________________________ by the number equivalent to the charge on the ion.

For negative ions this _________________________________ by the number equivalent to the charge on the ion.

Worksheet 2: Orbitals and electron configuration

Fill in the following table.

Quantum shell	Maximum number of electrons	Types of orbitals	Total number of orbitals	Electron configuration
$n = 1$				
$n = 2$				
$n = 3$				
$n = 4$				

Sketch the shapes of the *s* and *p* orbitals.

s

p

Complete the following table to show the electron configuration of the elements in the first column.

		Electron configuration		Electrons in boxes														
	Z	2.8.8	s, p, d	__s	__s	__p			__s	__p			__d					__s
Na																		
Be																		
Be^{2+}																		
P																		
Cr																		
Cu																		
Fe																		
Al																		
Al^{3+}																		
Sc																		
Cl																		
Cl^-																		

Examples of students’ responses from Results Plus – Examiners’ report

Here are some examples of answers – you may want to print out the answers and ask your students to mark them before sharing the examiners’ commentaries.

Example 1

15 The relative atomic mass of an element is determined using a mass spectrometer.

(a) Define the term **relative atomic mass**.

(2)

Relative atomic mass is the mass of an atom of an element relative to the mass of 1/12 of the atom of carbon 12.

First mark is NOT awarded as no mention of average/mean.
Second mark awarded as mention of carbon-12.

Example 2 – representations of dot-and-cross diagrams

Perfect answer:
1. Correct charge on BOTH ions.
2. Correct number of outer electrons.
3. No overlap of electron shells – clear separation of ions.

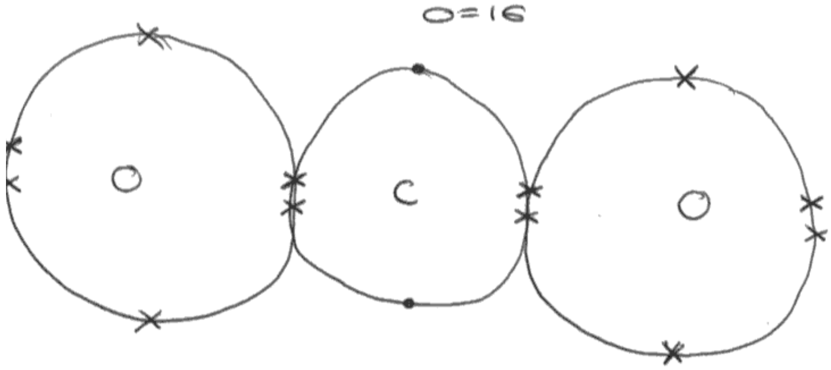
Example 3

(iii) Draw a dot and cross diagram of a molecule of carbon dioxide.

Show outer electrons only.

C = 12
O = 16

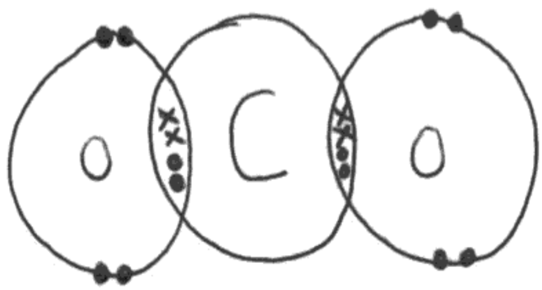
(2)



Wrong number of outer electrons for all the atoms shown.

The covalent bonds shown represent electrons donated by the oxygen.

No marks awarded.



Full marks – note that the total number of electrons on each atom's outer shell is 8.

Example 4

21 (a) Define the term **relative isotopic mass**.

(2)

The weighted average of all the masses of the isotopes of an element relative to 1/12 of carbon-12 atom.

The first mark was not awarded as the plural (i.e. isotopes) has been used and confusion is evident with definition of relative atomic mass.

The second mark is awarded as carbon-12 is mentioned.

Example 5

(ii) Explain what is meant by the term **isotopes**.

(2)

Isotopes are different forms of one element

(ii) Explain what is meant by the term **isotopes**.

(2)

Isotopes are diggerent atomic structures of the same element, with the same number of protons but diggerent number of neutrons.

The second answer is a good answer with both points given – same number of protons and different number of neutrons.

The candidate has indicated that they are atoms of the SAME element.

(b) Each element has an atomic number.

(i) State what is meant by **atomic number**.

(1)

Atomic number is the number of protons and electrons of an element.

(b) Each element has an atomic number.

(i) State what is meant by **atomic number**.

(1)

The + total ~~number~~ number of protons and neutrons there are in an atom.

First answer – may be correct but is unclear (total number or either number?).

Second answer – candidate has confused this with mass number.

Exam practice

1 The relative atomic mass of an element is determined using a mass spectrometer. State what is meant by the term *relative atomic mass*.

(2 marks)

(Edexcel GCE Jan 2011, 6CH01, Q15a)

2 Chlorine forms compounds with magnesium and with carbon.

a Draw a dot-and-cross diagram to show the electronic structure of the compound magnesium chloride (only the outer electrons need be shown). Include the charges present.

(2 marks)

b Draw a dot-and-cross diagram to show the electronic structure of the compound tetrachloromethane (only the outer electrons need be shown).

(2 marks)

(Edexcel GCE Jan 2011, 6CH01, Q17b)

c Draw a dot-and-cross diagram of a molecule of carbon dioxide. Show outer electrons only.

(2 marks)

(Edexcel GCSE Jun 2013, 5CH2H, Q2(iii))

3 a State what is meant by the term *relative isotopic mass*.

(2 marks)

(Edexcel GCE Jun 2012, 6CH01, Q21(a))

b State what is meant by the term *isotopes*.

(2 marks)

(Edexcel GCSE May 2012, 5CH2H, Q2bii)

c i State what is meant by the term *relative atomic mass*.

(2 marks)

ii A sample of boron contains:

- 19.7% of boron-10
- 80.3% of boron-11.

Use this information to calculate the relative atomic mass of boron.

(3 marks)

(Edexcel GCSE May 2013, 5CH2H, Q4c(i)–(ii))

4 A molecule is …

- **A** a group of atoms joined by ionic bonding.
- **B** a group of atoms joined by covalent bonding.
- **C** a group of ions joined by covalent bonding.
- **D** a group of atoms joined by metallic bonding.

(1 mark)

5 The relative atomic mass is defined as …

- **A** the mass of an atom of an element relative to $\frac{1}{12}$ the mass of a carbon-12 atom.
- **B** the mass of an atom of an element relative to the mass of a hydrogen atom.
- **C** the average mass of an element relative to $\frac{1}{12}$ the mass of a carbon atom.
- **D** the average mass of an atom of an element relative to $\frac{1}{12}$ the mass of a carbon-12 atom.

(1 mark)

(Edexcel GCE Jan 2012, 6CH01, Q1,2)

6 Which pair of ions is isoelectronic?

- **A** Ca^{2+} and O^{2-}
- **B** Na^{+} and O^{2-}
- **C** Li^{+} and Cl^{-}
- **D** Mg^{2+} and Cl^{-}

(1 mark)

(Edexcel GCE May 2013, 6CH01, Q4)

7 The isotopes of magnesium ${}^{24}_{12}Mg$ and ${}^{25}_{12}Mg$ both form ions with charge 2+. Which of the following statements about these ions is true?

- **A** Both ions have electronic configuration $1s^2\ 2s^2\ 2p^6\ 3s^2$.
- **B** ${}^{25}_{12}Mg^{2+}$ has more protons than ${}^{24}_{12}Mg^{2+}$.
- **C** The ions have the same number of electrons but different numbers of neutrons.
- **D** The ions have the same number of neutrons but different numbers of protons.

(1 mark)

8 Chlorine has two isotopes with relative isotopic mass 35 and 37. Four m/z values are given below. Which will occur in a mass spectrum of chlorine gas, Cl_2, from an ion with a single positive charge?

- **A** 35.5
- **B** 36
- **C** 71
- **D** 72

(1 mark)

(Edexcel GCE Jan 2010, 6CH01, Q1,2)

9 The electronic structures of four elements are given below. Which of these elements has the highest first ionisation energy?

	1*s*	2*s*	2*p*		
A	↑↓	↑↓	↑	↑	
B	↑↓	↑↓	↑	↑	↑
C	↑↓	↑↓	↑↓	↑↓	↑
D	↑↓	↑↓	↑↓	↑↓	↑↓

(1 mark)

10 Which of the following represents the electronic structure of a nitrogen atom?

	1*s*	2*s*	2*p*		
A	↑↓	↑	↑↓	↑	↑
B	↑↓	↑	↑↓	↑↓	
C	↑↓	↑↓	↑	↑	↑
D	↑↓	↑↓	↑↓	↑	

(1 mark)

(Edexcel GCE May 2012, 6CH01, Q19,20)

11 a In atoms, electrons fill up the sub-shells in order of increasing energy.

Complete the outer electronic configuration for an arsenic and a selenium atom using the electrons-in-boxes notation.

(2 marks)

		4s	4p		
As	[Ar] $3d^{10}$	↑↓			
Se	[Ar] $3d^{10}$	↑↓			

(Edexcel GCE Jan 2010, 6CH01, Q9c)

b Electrons in atoms occupy orbitals.

i Explain the term *orbital*.

(1 mark)

ii Draw diagrams below to show the shape of an *s*-orbital and of a *p*-orbital.

s-orbital	*p*-orbital

(2 marks)

c State the total number of electrons occupying all the *p*-orbitals in one atom of chlorine.

(1 mark)

d State the number of electrons present in an ion of calcium, Ca^{2+}.

(1 mark)

(Edexcel GCE May 2013, 6CH01R Q21)

12 The following data were obtained from the mass spectrum of a sample of platinum.

Peak at *m/z*	%
194	32.8
195	30.6
196	25.4
198	11.2

a Calculate the relative atomic mass of platinum in this sample. Give your answer to one decimal place.

(2 marks)

b In which block of the Periodic Table is platinum found?

(1 mark)

(Edexcel GCE May 2013, 6CH01R, Q21,22b,c)

13 The radioactive isotope iodine-131, $^{131}_{53}I$, is formed in nuclear reactors providing nuclear power. Naturally occurring iodine contains only the isotope $^{127}_{53}I$.

a Complete the table to show the number of protons and neutrons in these two isotopes.

Isotope	$^{131}_{53}I$	$^{127}_{53}I$
Number of protons		
Number of neutrons		

(2 marks)

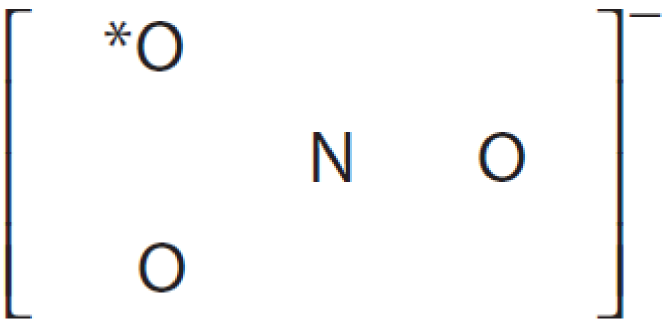
b When iodine-131 decays, one of its neutrons emits an electron and forms a proton. Identify the new element formed.

(1 mark)

(Edexcel GCE May 2013, 6CH01R, 18a,b)

14 The nitrate ion, NO_3^-, contains both covalent and dative covalent bonds. Complete the dot-and-cross diagram to show the bonding in the nitrate ion.

Only the outer electron shells for each atom need to be shown.
Represent the nitrogen electrons with crosses (×), and oxygen electrons with dots (•).
The symbol * on the diagram represents the extra electron giving the ion its charge.



(3 marks)

(Edexcel GCE May 2014R, 6CH01R, 20d)

Section B: Quantitative analysis and equations

This section covers one of the most important areas of the chemistry specification. A good understanding of the concepts covered here, particularly reacting masses, will have a huge impact on students' studies of later topics, including the A2 specification. The Table below lists the areas that students most commonly struggle with.

Perhaps the biggest barrier is understanding what is being asked when a practical scenario is given. We have provided a worked example of such questions. Unlike Physics, formulae and equations are not provided in Chemistry exams so it is important that students know these very well and, more importantly, be able to manipulate them as necessary to solve a given problem.

Students' strengths and common misconceptions

The table below outlines the areas in which most students do well and the common mistakes and misconceptions across the topics listed.

	Strengths	**Common mistakes**
Quantities in chemistry	Definitions as 'standalone'.	Conversions from one quantity to another, e.g. moles to grams. Not recognising that molar quantities are the same but the method of calculation depends on the species e.g. solutes in solution, gases, solids.
Empirical formulae	Writing empirical formula from molecular formula. Recognising a mathematical relationship between % composition and A_r.	Inverting the %/A_r ratio. Failing to simplify the ratios. Writing a final answer. Deducing molecular formula from empirical formula and M_r.
Balancing equations	Simple acid–alkali and metal plus oxygen or halogen equations.	Translating practical scenarios into word and formula equations. Not learning the common 'known' reactions e.g. carbonate plus acid. Applying the law of conservation of mass to equations. Balancing equations with diprotic acids.
Ionic equations	Given the state symbols, be able to split the ions.	Not knowing which species are soluble and the state symbols of common chemical species. Splitting common acids.
Reacting masses	Conservation of mass. Working out masses or moles as standalone direct questions.	Selecting the correct formula when solving problems with practical scenarios. Following multistep procedures and calculations.

Table of resources in this section

Topics covered	Type of resource	Resource name	Brief description and notes for resource
• Isotopes • Equations • Reacting masses	Teacher resource	Building knowledge	Building knowledge learning outcomes. May be used to assess understanding and for reflection on learning. Used for setting targets for improvement.
• Writing formulae • Reacting masses • Percentage yield	Teacher resource	Summary sheet: Writing formulae	Review of KS4 concepts. Summary of key points and guide to correct use of key terms. Tips on how to answer exam questions.
• Empirical formulae • Molar volumes • Avogadro constant	Teacher and student resource	Worked examples: Calculations	
• Writing chemical formulae	Student worksheet	Worksheet 1: Chemical formulae Worksheet 2: Cations and anions Worksheet 3: Writing equations	Practice working out molecular formulae from names of compounds. Checking understanding of new KS5 learning.
• Quantitative analysis and calculations	Student questions	Exam practice	How to answer exam-type questions and KS5 level. Covering main misconceptions for main topics.

Building knowledge

Quantitative chemistry

Isotopes – Why is the A_r of some elements not a whole number?

Building your understanding

Deduce the % abundance of a given isotope from data of the other isotopes and A_r.

Calculate the relative atomic mass of an element given the % abundances of its isotopes.

Give the similarities and differences between atoms of the same element (definition of isotopes).

Research task: *how do we investigate the presence of isotopes and their relative abundances?*

Quantitative chemistry

Equations and reacting masses

Building your understanding

- Given a reaction in words, write a balanced symbol equation.
- Write down ionic equations and know which ions can be omitted.
- Write equations with state symbols for chemical reactions from observations recorded.
- Know how to balance equations.
- Deduce the formulae for compounds with more complex anions (compound ions).
- Deduce the formulae for simple ionic compounds with just two types of elements.
- Work out the charge on an ion from its position in the Periodic Table.

Summary sheet: Writing formulae

Writing formulae

Compounds should have no overall charges, so the positive and negative charges should cancel each other out.

Apart from working out the charges on ions made up of one element, you need to know the following compound ions and their charges.

Name	Formula	Charge
hydroxide	OH^-	1−
nitrate	NO_3^-	1−
sulfate	SO_4^{2-}	2−
carbonate	CO_3^{2-}	2−
ammonium	NH_4^+	1+

Follow these steps.

Write the name of the compound	Magnesium bromide	Sodium sulfate
Work out the charge of your positive ion = group number, or 1+ for ammonium.	Mg^{2+}	Na^+
Work out the charge of your negative ion = group number − 8 *or* known charge for a compound ion.	Br^-	SO_4^{2-}
Rewrite the symbols; put a bracket around any compound ion.	Mg^{2+} Br^- Mg Br	Na^+ SO_4^{2-} Na (SO_4)
Swap the numbers of the charges and drop them to the opposite ion.	$MgBr_2$	$Na_2(SO_4)$

Writing ionic equations

- Make sure all state symbols are included.
- Identify the species that are aqueous, using the rules of solubility.

1 Look at the cation – is it Group 1 or ammonium? If so → soluble.

2 Look at the anion – is it a nitrate? If so → soluble.

- Proceed only if you have ruled out 1 and 2.

1 Is the anion a halide (chloride, bromide or iodide)?

2 If so, look at the metal – lead or silver? If so → insoluble.

3 Is the anion a sulfate?

4 If so, look at the metal – barium, calcium, lead? If so → insoluble.

5 Is the anion a hydroxide?

6 If so, look at the metal – transition metal or Group 2 (after Ca)? If so → insoluble.

- Split all the soluble salts into their aqueous ions on both sides – remember to write the numbers in front of the ions for multiples.
- Cancel out the ions that appear on both sides – again pay attention to numbers.
- Write your final equation (always keep the state symbols unless specifically told not to!).

Reacting masses

To work out masses of reactants and products from equations, follow these steps.

Steps to follow	Example	Example
	5 g of Ca reacted with excess chlorine. What mass of $CaCl_2$ is formed?	When $MgCO_3$ was heated strongly, 4 g of MgO was formed. What is the mass of $MgCO_3$ that was heated?
Write the balanced equation.	$Ca + Cl_2 \rightarrow CaCl_2$	$MgCO_3 \rightarrow MgO + CO_{2(g)}$
Write the masses given.	5 g (excess) **?**	**?** 4 g
Find the A_r or M_r.	40 111	84 40
Divide by the atomic or molecular mass (step 2 ÷ step 3).	$\frac{5}{40} : \frac{?}{111}$	$\frac{?}{84} : \frac{4}{40}$
Treat these like ratios, rearrange to find the unknown (**?**).	Mass of $CaCl_2$ = (5 × 111) ÷ 40 = 13.9 g	Mass of $MgCO_3$ = (4 × 84) ÷ 40 = 8.4 g

Note: if you are told something is in excess, do not use it in the calculation!

Percentage yield

The calculations above dealt with the masses you get or use if the reaction is 100% complete.

Most reactions are not 100% complete for the following reasons:

- not all the reactant reacts
- some is lost in the glassware as you transfer the reactants and the products
- some other products might be formed that you do not want.

This is a problem in industry. Less of the desired product has been made, so there is less to use or sell, and the waste has to be disposed of. Waste products can be harmful to the environment, e.g. the one above produces the greenhouse gas CO_2. Industries try to choose reactions that minimise waste and do not produce harmful products. They also try to make the rate of reaction high enough to make the reaction turnover fast so they can increase production and make money.

To work out % yield: use the balanced equation to work out how much of the given product you should get if the reaction is 100% efficient – this is the theoretical yield.

Then: $\%\text{ yield} = \frac{\text{actual yield} \times 100}{\text{theoretical yield}}$

Worked examples: Calculations

The example exam questions in the shaded sections are followed by working out and hints on answering the questions.

Empirical formulae

1 Sulfamic acid is a white solid used by plumbers as a limescale remover.

a Sulfamic acid contains 14.42% by mass of nitrogen, 3.09% hydrogen and 33.06% sulfur. The remainder is oxygen.

i Calculate the empirical formula of sulfamic acid. **(3)**

Interpreting the question

- 'The remainder is oxygen.' So you need to calculate the percentage of oxygen.
- 'Calculate the empirical formula of sulfamic acid.' This is the main question.

Answering the question

What you do	Calculation				Common mistakes
Write the symbols of the elements.	N	H	S	O	Remember you can check the symbols in the Periodic Table.
Note the % underneath.	14.42	3.09	33.06	100 – (14.42 + 3.09 + 33.06) = 49.43	Check sum of % = 100%. Make sure you transfer the correct % for the correct element.
Write the A_r.	14.01	1	32.06	16	Remember to use the Periodic Table correctly!
Divide % by A_r for ratio.	1.03	3.09	1.03	3.09	Do not round up at this stage.
Divide by smallest number for simplest ratio.	1	3	1	3	These numbers give you the number of each atom in the empirical formula.
Write the empirical formula.	NH_3SO_3				Make sure you actually write this formula out – don't leave the answer at the ratio stage.

ii The molar mass of sulfamic acid is 97.1 g mol^{-1}. Use this information to deduce the molecular formula of sulfamic acid.

Answering the question

Work out empirical mass first, then use this to work out the molecular formula.

1 1 × N = 14; 3 × H= 3; 1 × S = 32; 3 × O = 16 × 3 = 48

2 Empirical mass =14 + 3 + 32 + 48 = 97

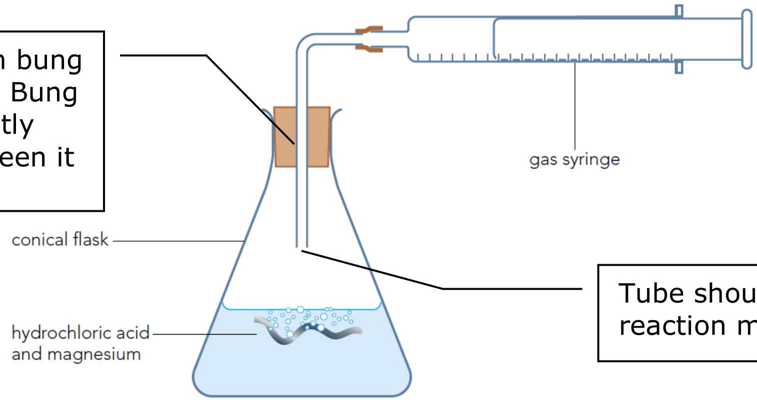
3 Divide molar mass by empirical mass: 97.01/97 = 1, therefore molecular formula = empirical formula.

b **S**ulfamic acid reacts with magnesium to produce hydrogen gas. In an experiment, a solution containing 5.5×10^{-3} moles of sulfamic acid reacted with excess magnesium. The volume of hydrogen produced was 66 cm^3, measured at room temperature and pressure.

i Draw a labelled diagram of the apparatus you would use to carry out this experiment, showing how you would collect the hydrogen produced and measure its volume.

Answering the question

Draw tube through bung as a cross section. Bung should also fit tightly with no gaps between it and the flask.



Tube should not touch reaction mixture.

ii Calculate the number of moles of hydrogen, H_2, produced in this reaction.

The molar volume of a gas is 24 $dm^3\,mol^{-1}$ at room temperature and pressure.

Interpreting the question

- *Excess magnesium* means that you cannot use this substance in the calculation.
- The molar volume is given in dm^3 but the volume of hydrogen is given in cm^3.

Answering the question

1 The molar volume of a gas is 24 $dm^3\,mol^{-1}$ at room temperature and pressure.

2 Number of moles of a gas = volume/molar volume.

3 Number of moles of H_2 = 66/24 000 = 2.75×10^{-3} mol.

iii Show that the data confirms that two moles of sulfamic acid produces one mole of hydrogen gas, and hence write an equation for the reaction between sulfamic acid and magnesium, using $H[H_2NSO_3]$ to represent the sulfamic acid.

Interpreting the question

This question is asking you to compare the number of moles.

- sulfamic acid = 5.5×10^{-3} mol.
- hydrogen molecules = 2.75×10^{-3} mol (answer from part **ii**).

Answering the question

1 5.5×10^{-3} mol of sulfamic acid produce 2.75×10^{-3} mol of H_2, so

2 2 mol of sulfamic acid produce 1 mol of H_2

3 $2\ H[H_2NSO_3] + Mg \rightarrow Mg(H_2NSO_3)_2 + H_2$

Molar gas volumes and the Avogadro constant

2 Airbags, used as safety features in cars, contain sodium azide, NaN_3. An airbag requires a large volume of gas produced in a few milliseconds. The gas is produced in this reaction:

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g) \qquad \Delta H \text{ is positive}$$

When the airbag is fully inflated, it contains 50 dm^3 of nitrogen gas.

a Calculate the number of molecules in 50 dm^3 of nitrogen gas under these conditions.

[The Avogadro constant = 6.02×10^{23} mol^{-1}. The molar volume of nitrogen gas under the conditions in the airbag is 24 $dm^3\ mol^{-1}$.]

Interpreting the question

- The Avogadro constant is used when you need to work out the number of particles.
- When you are given the molar volume, you will need to calculate the number of moles.

Answering the question

1 Use molar volume to convert 50 dm^3 to moles of N_2.

Number of moles of N_2 = 50/24 = 2.08 mol

2 Use the Avogadro constant to work out the number of molecules in 2.08 mol.

$6.02 \times 10^{23} \times 2.08 = 1.25 \times 10^{24}$ molecules

b Calculate the mass of sodium azide, NaN_3, that would produce 50 dm^3 of nitrogen gas.

Answering the question

1 Molar ratios: $2NaN_3 \rightarrow 2Na + 3N_2$

2 Number of moles: ? ? 2.08

The question asks us to relate sodium azide to nitrogen gas. Using the equation, you can see that every 2 mol of sodium azide (NaN_3) gives 3 mol of nitrogen (N_2). Therefore the number of moles of sodium azide is always two-thirds that of nitrogen.

3 Using ratios: number of moles of sodium azide = $\frac{2}{3} \times 2.08 = 1.39$ mol.

4 Convert moles to mass:

- Molar mass of sodium azide = $23 + (14 \times 3) = 65\,g\,mol^{-1}$
- Use equation: Number of moles = mass/molar mass
  so mass = number of moles × molar mass = $65\,g\,mol^{-1} \times 1.39\,mol = 90.4\,g$

Worksheet 1: Chemical formulae

Write the formulae of the following compounds.

Copper(II) sulfate	____________________
Nitric acid	____________________
Copper(II) nitrate	____________________
Sulfuric acid	____________________
Sodium carbonate	____________________
Aluminium sulfate	____________________
Ammonium nitrate	____________________
Nitrogen dioxide	____________________
Sulfur dioxide	____________________
Ammonia	____________________
Ammonium sulfate	____________________
Potassium hydroxide	____________________
Calcium hydroxide	____________________

Worksheet 2: Cations and anions

Complete the table below to show the substance, its formula and its individual ions.

Substance	**Formula**	**Cation (exact number)**	**Anion (exact number)**
Sodium bromide			
	KI		
Silver nitrate			
Copper(II) sulfate			
	$NaHCO_3$		
Magnesium carbonate			
Lithium carbonate			
	$Ca(HSO_4)_2$		
Aluminium nitrate			
Calcium phosphate			
Potassium hydride			
Sodium ethanoate			
	$KMnO_4$		
Potassium dichromate(VI)			
Zinc chloride			
Strontium nitrate			
Sodium chromate(VI)			
Calcium fluoride			
Potassium sulfide			
Magnesium nitride			
Lithium hydrogensulfate			
	$(NH_4)_2SO_4$		

Worksheet 3: Writing equations

Write: **(a) the chemical equation and (b) the ionic equation for each of the following reactions.**

1 Magnesium with sulfuric acid

2 Calcium carbonate with nitric acid

3 Hydrochloric acid with sodium hydroxide

4 Aqueous barium chloride with aqueous sodium sulfate

5 Aqueous sodium hydroxide with sulfuric acid

6 Aqueous silver nitrate with aqueous magnesium chloride

7 Solid magnesium oxide with nitric acid

8 Aqueous copper(II) sulfate with aqueous sodium hydroxide

9 Aqueous lead(II) nitrate with aqueous potassium iodide

10 Aqueous iron(III) nitrate with aqueous sodium hydroxide

Exam practice

1 Coral reefs are produced by living organisms and predominantly made up of calcium carbonate. It has been suggested that coral reefs will be damaged by global warming because of the increased acidity of the oceans due to higher concentrations of carbon dioxide.

a Write a chemical equation to show how the presence of carbon dioxide in water results in the formation of carbonic acid. State symbols are not required.

(1 mark)

b Write the ionic equation to show how acids react with carbonates. State symbols are not required.

(2 marks)

2 One method of determining the proportion of calcium carbonate in a coral is to dissolve a known mass of the coral in excess acid and measure the volume of carbon dioxide formed.

In such an experiment, 1.13 g of coral was dissolved in 25 cm^3 of hydrochloric acid (an excess) in a conical flask. When the reaction was complete, 224 cm^3 of carbon dioxide had been collected over water using a 250 cm^3 measuring cylinder.

a Draw a labelled diagram of the apparatus that could be used to carry out this experiment.

(2 marks)

b Suggest how you would mix the acid and the coral to ensure that no carbon dioxide escaped from the apparatus.

(1 mark)

c Calculate the number of moles of carbon dioxide collected in the experiment. (The molar volume of any gas is 24 000 $cm^3 mol^{-1}$ at room temperature and pressure.)

(1 mark)

d Complete the equation below for the reaction between calcium carbonate and hydrochloric acid by inserting the missing state symbols.

$CaCO_3$(........) + 2HCl(........) → $CaCl_2$(........) + H_2O(l) + CO_2(........)

(1 mark)

e Calculate the mass of 1 mol of calcium carbonate. (Assume relative atomic masses: Ca = 40.1, C = 12.0, O = 16.0)

(1 mark)

f Use your data and the equation in **d** to calculate the mass of calcium carbonate in the sample and the percentage by mass of calcium carbonate in the coral. Give your final answer to three significant figures.

(2 marks)

g When this experiment is repeated, the results are inconsistent. Suggest a reason for this other than errors in the procedure, measurements or calculations.

(1 mark)

3 Magnesium chloride can be made by reacting solid magnesium carbonate, $MgCO_3$, with dilute hydrochloric acid.

a Write an equation for the reaction, including state symbols.

(2 marks)

b A precipitate of barium sulfate is produced when aqueous sodium sulfate is added to aqueous barium chloride. Give the ionic equation for the reaction, including state symbols.

(2 marks)

Section C: Structure and properties – Literacy Focus

In this section we apply the concepts covered in Section A to properties of materials. The resources provided highlight the importance of selecting the correct key words when describing and explaining properties of materials. One of the most effective ways of helping students construct extended writing is by using key word maps, where they are asked to select the appropriate key words from a list. Part of their learning is the ability to select the correct terms needed for a given task.

The teacher resources give the learning outcomes, and the summary sheets look back at what was taught at KS4. As teachers we are very good at telling students what to write in exams but not what they should not write. Therefore we have focused on this area in all three sections, most importantly in this section, which aims to help students improve their scientific writing. We envisage that they will understand that terms like *molecules* and *ions* are not interchangeable and they will learn to be more selective and specific with the scientific terms they use.

Students' strengths and common misconceptions

The table below outlines the general areas in which students do well and the common mistakes and misconceptions across the topics listed.

	What most students can do (well)	**Common mistakes**
Metals	Stating the physical properties of metals, including conductivity. Describing the structure as particles with delocalised electrons.	Using words like *molecules* and *atoms* instead of *cations* or *ions*, and *free* instead of *delocalised* or *free-moving* electrons to describe metallic bonds. Explaining the differences in the melting point and electrical conductivity of two metals.
Ionic compounds	Knowing that ionic compounds form giant structures, and therefore have high melting points. Knowing that ionic compounds conduct electricity when molten or in solutions.	In explaining or describing the electrostatic attraction between cations and anions in the giant structure. When describing separation of the ions at melting temperature. Explaining why ionic compounds conduct electricity when molten or in solution using terms like *free electrons* instead of in terms of mobility of ions.
Covalent compounds	Knowing the existence of simple molecular and giant covalent structures and give examples of each. Knowing of the existence of intermolecular forces and the effect of increasing molecular mass. In diamond each carbon atom forms covalent bonds with four others whereas in graphite it bonds only with three.	Explaining the boiling point – distinguishing between intermolecular forces in simple molecules and extensive covalent bonds in giant structures.

Table of resources in this section

Topics covered	Type of resource	Resource name	Brief description and notes for resource
• Ionic bonding	Teacher resource	Building knowledge	Building knowledge learning outcomes. May be used to assess understanding and for reflection on learning. Used for setting targets for improvement.
• Ionic bonding	Teacher resource	Summary sheets	Selecting the correct vocabulary to describe bonding and properties of ionic compounds, metals and covalent compounds.
• Ionic structure	Teacher resource	Teaching ideas: Key words to describe ionic structure	Literacy activity – scaffolding resource: • how to structure long descriptive answers using the correct key words • relating physical properties to bonding.
	Student questions	Exam practice	Using skills learnt in this section to write succinct answers using precise vocabulary.

Teacher resources

The big questions

- What does a material need to have in order to conduct electricity?
- When do ionic compounds conduct electricity?
- How can this be explained in terms of the nature of the bonds?

Building knowledge

Ionic bonding

Structure and properties

Building your understanding

Deduce the type of structure of a substance from solubility and electrical conductivity data.

Relate the physical properties of ionic compounds to their bonding and structure.

Explain why most ionic compounds are soluble in water.

Describe the lattice structure of ionic compounds.

Summary sheet 1: Structure and bonding

Words used to describe structure and bonding:

- ions, atoms, molecules, intermolecular forces, electrostatic forces, delocalised electrons, cations, anions, outer electrons, shielding

Metallic bond: electrostatic attraction between the nuclei of cations (positive ions) and delocalised electrons.

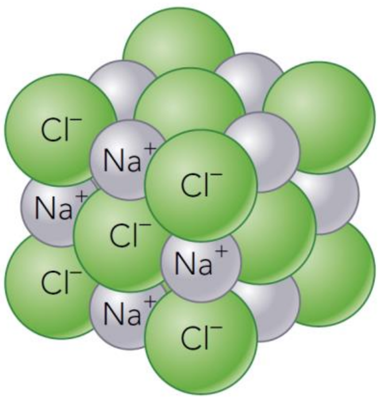
Strength of the metallic bonding increases with the number of valence electrons (outer electrons in the atoms) and with decreasing size of the cation.

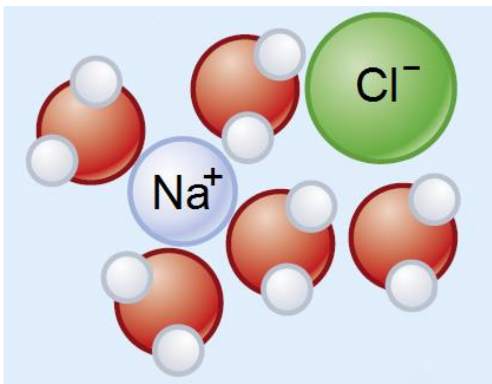
Ionic bonds and ionic compounds

> Explain why NaCl has a high melting point and only conducts electricity when molten or in solution. (6 marks)

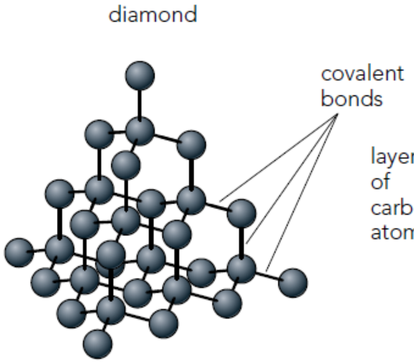
An answer should cover the following points.

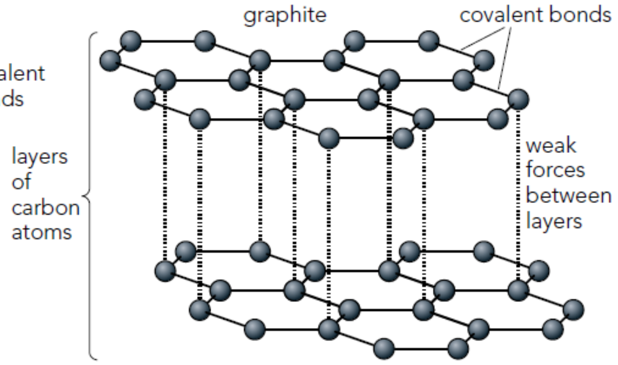
1. The Na^+ and Cl^- ions are held by strong electrostatic forces.
2. To melt solid NaCl, energy is needed to separate overcome the forces of attraction sufficiently for the lattice structure to break down and for the ions to be free to slide past one another.
3. Even though the ions are charged, the solid cannot conduct electricity because the ions are not mobile (free to move).
4. If the solid is melted, the ions can move freely and allow the liquid to conduct electricity.
5. Also, when dissolved in water the *ions* are separated by the water molecules and so are free to move, hence the aqueous solution can conduct electricity.





Summary sheet 2: Diamond and graphite structures



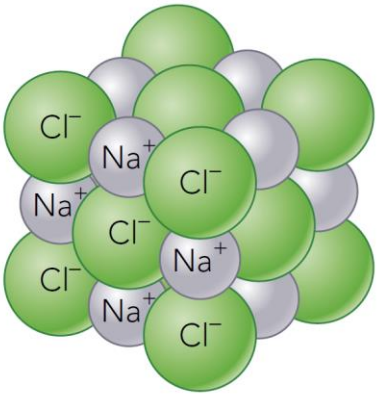


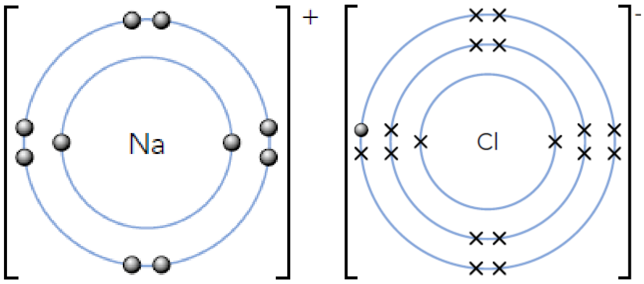
Property	Diamond	Graphite
Melting point	High – atoms held by strong covalent bonds. Many covalent bonds must be broken to melt it. Is solid at room temp.	High – atoms held by strong covalent bonds. Many covalent bonds must be broken to melt it. Is a solid at room temp.
Electrical conductivity	Poor – no mobile electrons available. All 4 outer electrons of each carbon are used in bonding.	Good – each carbon only uses 3 of its outer electron to form covalent bonds. 4^{th} electron form each atom contributes to a delocalised electron system. These delocalised electrons can flow when a potential difference is applied parallel to the layers.
Lubricant	Poor – structure is rigid.	Gas molecules are trapped between the layers and allow the layers to slide past one another. Same reason for its use in pencils.
Solubility	Insoluble in water – no charged particles to interact with water (think of SiO_2, main component of sand).	Insoluble in water – no charged particles to interact with water (think of SiO_2, main component of sand).

Teaching ideas: Using key words to describe ionic structure

Describe and explain how the structure of sodium fluoride is formed.

Use knowledge of the structure of sodium chloride





Which key words will you need?

- Attraction
- Electrostatic
- Tight
- Non-metals
- Giant
- Packed
- Anions
- Strong
- Metals
- Forces
- Ionic
- Opposition
- Lattice
- Cations

Tip

For questions about the physical properties of ionic compounds, relate the properties to their bonding and structure.

Property	**Why?**
Does not conduct electricity when solid.	
Conducts electricity when molten or in aqueous solution.	
	The ions are held by strong electrostatic forces of attraction and a large amount of energy is needed to overcome the attractions.
	The ions are tightly packed together.

Exam practice

1 Suggest why the melting temperature of magnesium oxide is higher than that of magnesium chloride, even though both are almost 100% ionic.

(3 marks)

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2 Silicon exists as a giant covalent lattice.

a The electrical conductivity of pure silicon is very low. Explain why this is so in terms of the bonding.

(2 marks)

b Explain the high melting temperature of silicon in terms of the bonding.

(2 marks)

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3 The melting temperatures of the elements of Period 3 are given in the table below. Use these values to answer the questions that follow.

Element	Na	Mg	Al	Si	P (white)	S (monoclinic)	Cl	Ar
Melting temperature / K	371	922	933	1683	317	392	172	84

a Explain why the melting temperature of sodium is very much less than that of magnesium.

(3 marks)

b Explain why the melting temperature of silicon is very much greater than that of white phosphorus.

(3 marks)

c Explain why the melting temperature of argon is the lowest of all the elements of Period 3.

(1 mark)

d Explain why magnesium is a good conductor of electricity whereas sulfur is a non-conductor.

(2 marks)

Appendices

Appendix 1: Specification mapping

Key

(orange)	5CH1F/H – Core Science
(purple)	5CH2F/H – Additional Science
(green)	5CH3F/H – Extension Unit or Further Additional Science

The table on the following pages maps certain topics from the new AS level Chemistry specification across to relevant sections within the GCSE specification.

Topic 1 – Atomic structure and the Periodic Table	GCSE
1. know the structure of an atom in terms of electrons, protons and neutrons	1.3 Describe the structure of an atom as a nucleus containing protons and neutrons, surrounded by electrons in shells (energy levels)

Topic 1 – Atomic structure and the Periodic Table	GCSE
2. know the relative mass and relative charge of protons, neutrons and electrons	1.4 Demonstrate an understanding that the nucleus of an atom is very small compared to the overall size of the atom 1.5 Describe atoms of a given element as having the same number of protons in the nucleus and that this number is unique to that element 1.6 Recall the relative charge and relative mass of: a a proton b a neutron c an electron 1.7 Demonstrate an understanding that atoms contain equal numbers of protons and electrons
3. know what is meant by the terms *atomic (proton) number* and *mass number*	1.8 Explain the meaning of the terms a atomic number b mass number
4. be able to determine the number of each type of sub-atomic particle in an atom, molecule or ion from the atomic (proton) number and mass number	1.9 Describe the arrangement of elements in the Periodic Table such that: a elements are arranged in order of increasing atomic number, in rows called periods b elements with similar properties are placed in the same vertical column, called groups
5. understand the term *isotopes*	1.10 Demonstrate an understanding that the existence of isotopes results in some relative atomic masses not being whole numbers

Topic 1 – Atomic structure and the Periodic Table	GCSE
6. be able to define the terms *relative isotopic mass* and *relative atomic mass*, based on the ^{12}C scale 7. understand the terms *relative molecular mass* and *relative formula mass*, including calculating these values from relative atomic masses Definitions of these terms will not be expected The term *relative formula mass* should be used for compounds with giant structures	1.8 State the meaning of the term c relative atomic mass
8. be able to analyse and interpret data from mass spectrometry to calculate relative atomic mass from relative abundance of isotopes and vice versa 9. be able to predict the mass spectra for diatomic molecules, including chlorine 10. understand how mass spectrometry can be used to determine the relative molecular mass of a molecule Limited to the *m*/*z* value for the molecular ion, M^+, giving the relative molecular mass of the molecule	2.16 Recall that chemists use spectroscopy (a type of flame test) to detect the presence of very small amounts of elements and that this led to the discovery of new elements, including rubidium and caesium 1.11 Calculate the relative atomic mass of an element from the relative masses and abundances of its isotopes

Topic 1 – Atomic structure and the Periodic Table	GCSE
16. know the number of electrons that can fill the first four quantum shells 17. know that an orbital is a region within an atom that can hold up to two electrons with opposite spins 18. know the shape of an s-orbital and a p-orbital 19. know the number of electrons that occupy s-, p- and d-sub-shells 20. be able to predict the electronic configurations, using 1s notation and electrons-in-boxes notation, of: i. atoms, given the atomic number, Z, up to Z = 36 ii. ions, given the atomic number, Z, and the ionic charge, for s- and p-block ions only, up to Z = 36 21. know that elements can be classified as s-, p- and d-block elements	1.12 Apply rules about the filling of electron shells (energy levels) to predict the electronic configurations of the first 20 elements in the Periodic Table as diagrams and in the form 2.8.1 1.13 Describe the connection between the number of outer electrons and the position of an element in the Periodic Table

Topic 2 – Bonding and structure	GCSE
1. know that ionic bonding is the strong electrostatic attraction between oppositely charged ions	2.7 Describe the structure of ionic compounds as a lattice structure: a consisting of a regular arrangement of ions b held together by strong electrostatic forces of attraction between oppositely charged ions

Topic 2 – Bonding and structure	GCSE
3. understand the formation of ions in terms of electron loss or gain 4. be able to draw electronic configuration diagrams of cations and anions using dot-and-cross diagrams	2.1 Demonstrate an understanding that atoms of different elements can combine to form compounds by the formation of new chemical bonds 2.2 Describe how ions are formed by the transfer of electrons 2.3 Describe an ion as an atom or group of atoms with a positive or negative charge 2.4 Describe the formation of sodium ions, Na^+, and chloride ions, Cl^-, and hence the formation of ions in other ionic compounds from their atoms, limited to compounds of elements in groups 1, 2, 6 and 7
7. know that a covalent bond is the strong electrostatic attraction between two nuclei and the shared pair of electrons between them	3.1 State that a covalent bond is formed when a pair of electrons is shared between two atoms 3.2 Recall that covalent bonding results in the formation of molecules
8. be able to draw dot-and-cross diagrams to show electrons in simple covalent molecules, including those with multiple bonds and dative covalent (coordinate) bonds	3.3 Explain the formation of simple molecular, covalent substances using dot-and-cross diagrams, including: a hydrogen b hydrogen chloride c water d methane e oxygen f carbon dioxide

Topic 2 – Bonding and structure	GCSE
22. know that metallic bonding is the strong electrostatic attraction between metal ions and the sea of delocalised electrons	4.2 Describe the structure of metals as a regular arrangement of positive ions surrounded by a sea of delocalised electrons 4.3 Describe and explain the properties of metals, limited to malleability and the ability to conduct electricity 4.4 Recall that most metals are transition metals and that their typical properties include: a high melting point b the formation of coloured compounds
23. know that giant lattices are present in: i. ionic solids (giant ionic lattices) ii. covalently bonded solids, such as diamond, graphite and silicon(IV) oxide (giant covalent lattices) iii. solid metals (giant metallic lattices) 25. know the different structures formed by carbon atoms, including graphite, diamond and graphene	3.6 Demonstrate an understanding of the differences between the properties of simple molecular covalent substances and those of giant covalent substances, including diamond and graphite 3.7 Explain why, although they are both forms of carbon and giant covalent substances, graphite is used to make electrodes and as a lubricant, whereas diamond is used in cutting tools
24. know that the structure of covalently bonded substances such as iodine, I_2, and ice, H_2O, is simple molecular 26. be able to predict the type of structure and bonding present in a substance from numerical data and/or other information	

Topic 2 – Bonding and structure	GCSE
27. be able to predict the physical properties of a substance, including melting and boiling temperature, electrical conductivity and solubility in water, in terms of: i. the types of particle present (atoms, molecules, ions, electrons) ii. the structure of the substance iii. the type of bonding and the presence of intermolecular forces, where relevant	3.4 Classify different types of elements and compounds by investigating their melting points and boiling points, solubility in water and electrical conductivity (as solids and in solution) including sodium chloride, magnesium sulfate, hexane, liquid paraffin, silicon(IV) oxide, copper sulfate, and sucrose (sugar) 3.5 Describe the properties of typical simple molecular, covalent compounds, limited to: a low melting points and boiling points, in terms of weak forces between molecules b poor conduction of electricity

Topic 5 – Formulae, equations and amounts of substance	GCSE
1. know that the mole (mol) is the unit for amount of a substance	6.1 Calculate relative formula mass given relative atomic masses 6.4 Calculate the percentage composition by mass of a compound from its formula and the relative atomic masses of its constituent elements
2. be able to use the Avogadro constant, *L* (6.02×10^{23} mol^{-1}), in calculations	2.1 Calculate the concentration of solutions in $g\,dm^{-3}$ 2.7 Demonstrate an understanding that the amount of a substance can be measured in grams, numbers of particles or number of moles of particles 2.8 Convert masses of substances into moles of particles of the substance and vice versa 2.9 Convert concentration in $g\,dm^{-3}$ into $mol\,dm^{-3}$ and vice versa

Topic 5 – Formulae, equations and amounts of substance	GCSE
3. know that the molar mass of a substance is the mass per mole of the substance in $g\,mol^{-1}$	
4. know what is meant by the terms *empirical formula* and *molecular formula*	6.2 Calculate the formulae of simple compounds from reacting masses and understand that these are empirical formulae 6.3 Determine the empirical formula of a simple compound, such as magnesium oxide
5. be able to calculate empirical and molecular formulae from experimental data Calculations of empirical formula may involve composition by mass or percentage composition by mass data	
6. be able to write balanced full and ionic equations, including state symbols, for chemical reactions	0.1 Recall the formulae of elements and simple compounds in the unit 0.2 Represent chemical reactions by word equations and simple balanced equations 0.3 Write balanced chemical equations including the use of state symbols (s), (l), (g) and (aq) for a wide range of reactions in this unit 0.4 Write balanced ionic equations for a wide range of reactions in this unit and those in unit C2, specification point 2.15
7. be able to calculate amounts of substances (in mol) in reactions involving mass, volume of gas, volume of solution and concentration These calculations may involve reactants and/or products	6.5 Use balanced equations to calculate masses of reactants and products

Topic 5 – Formulae, equations and amounts of substance	**GCSE**
8. be able to calculate reacting masses from chemical equations, and vice versa, using the concepts of amount of substance and molar mass	
9. be able to calculate reacting volumes of gases from chemical equations, and vice versa, using the concepts of amount of substance 10. be able to calculate reacting volumes of gases from chemical equations, and vice versa, using the concepts of molar volume of gases CORE PRACTICAL 1: Measure the molar volume of a gas	4.1 Demonstrate an understanding that one mole of any gas occupies $24\,dm^3$ at room temperature and atmospheric pressure and that this is known as the molar volume of the gas 4.2 Use molar volume and balanced equations in calculations involving the masses of solids and volumes of gases 4.3 Use Avogadro's law to calculate volumes of gases involved in gaseous reactions, given the relevant equations

Topic 5 – Formulae, equations and amounts of substance	GCSE
11. be able to calculate solution concentrations, in mol dm^{-3} and g dm^{-3}, for simple acid–base titrations using a range of acids, alkalis and indicators The use of both phenolphthalein and methyl orange as indicators will be expected CORE PRACTICAL 2: Prepare a standard solution from a solid acid and use it to find the concentration of a solution of sodium hydroxide CORE PRACTICAL 3: Find the concentration of a solution of hydrochloric acid 12. be able to: i. calculate measurement uncertainties and measurement errors in experimental results ii. comment on sources of error in experimental procedures 13. understand how to minimise the percentage error and percentage uncertainty in experiments involving measurements	2.12 Describe an acid-base titration as a neutralisation reaction where hydrogen ions (H^+) from the acid react with hydroxide ions (OH^-) from the base 2.13 Describe how to carry out simple acid–base titrations using burette, pipette and suitable acid–base indicators 2.14 Carry out an acid–base titration to prepare a salt from a soluble base 2.15 Carry out simple calculations using the results of titrations to calculate an unknown concentration of a solution or an unknown volume of solution required
14. be able to calculate percentage yields and percentage atom economies using chemical equations and experimental results Atom economy of a reaction = (molar mass of the desired product)/(sum of the molar masses of all products) × 100%	6.6 Recall that the yield of a reaction is the mass of product obtained in the reaction 6.7 Demonstrate an understanding that the actual yield of a reaction is usually less than the yield calculated using the chemical equation (theoretical yield) 6.8 Calculate the percentage yield

Topic 5 – Formulae, equations and amounts of substance	GCSE
15. be able to relate ionic and full equations, with state symbols, to observations from simple test tube reactions, to include: i. displacement reactions ii. reactions of acids iii. precipitation reactions 16. understand risks and hazards in practical procedures and suggest appropriate precautions where necessary.	2.13 Use solubility rules to predict whether a precipitate is formed when named solutions are mixed together and to name the precipitate 2.15 Describe tests to show the following ions are present in solids or solutions: b CO_3^{2-} using dilute acid and identifying the carbon dioxide evolved c SO_4^{2-} using dilute hydrochloric acid and barium chloride solution d Cl^- using dilute nitric acid and silver nitrate solution
	3.4 Recall that acids are neutralised by: a metal oxides b metal hydroxides c metal carbonates to produce salts (no details of salt preparation techniques or ions are required) 3.5 Recall that: a hydrochloric acid produces chloride salts b nitric acid produces nitrate salts c sulfuric acid produces sulfate salts

Appendix 2: Further baseline assessment questions

Section A: baseline assessment extra questions

1 Complete the table below.

	Number of protons	Number of electrons	Number of neutrons	Electron configuration
S				
Mg				
O^{2-}				
H^{+}				
Kr				
Al^{3+}				

(6 marks)

2 Draw a dot-and-cross diagrams for the following compounds.

a Methane

b Water

c Sodium fluoride

d Magnesium bromide

e Ammonia

f Potassium oxide

g Calcium oxide

h Oxygen

i Carbon dioxide

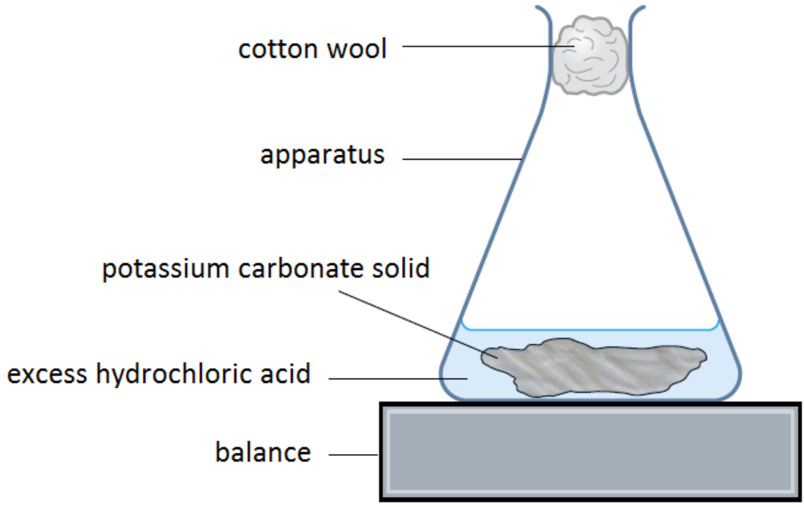
(18 marks)

Section B: baseline assessment extra questions

1 Magnesium has three isotopes. The mass spectrum of magnesium shows peaks at *m*/*z* 24 (78.60%), 25 (10.11%) and 26 (11.29%). Calculate the relative atomic mass of magnesium to 4 significant figures.

(2 marks)

2 2.76 g of solid potassium carbonate was reacted with excess hydrochloric acid, and the change in mass was recorded as shown in the diagram below.



The equation for the reaction is given by:

$$K_2CO_{3(s)} + 2HCl_{(aq)} \rightarrow 2KCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$$

The results from the experiment are:

- mass of K_2CO_3 + conical flask + HCl at the start = 194.05 g
- mass recorded at the end of the reaction = 193.39 g.

a Write the ionic equation for this reaction.

b Calculate the relative molecular mass M_r of K_2CO_3.

c Calculate the maximum mass of carbon dioxide which should be produced.

d Deduce the mass of carbon dioxide produced, and hence work out the % yield.

e What is the purpose of the cotton wool?

f Give two possible reasons why the yield is not 100%.

(9 marks)

Section C: baseline assessment extra questions

1 Complete the Table below using the following words:

ionic covalent giant simple metallic

Substance	Formula	Type of bonding	Structure
Hydrogen sulfide			
Graphite			
Silicon dioxide			
Calcium			
Magnesium chloride			
Fluorine			
Argon			

(7 marks)

2 By considering the type of bonding and structure, explain why aluminium melts at a higher temperature than lithium.

(3 marks)

3 Explain why potassium chloride does not conduct electricity when solid whereas copper does

(3 marks)