

INTERNATIONAL ADVANCED LEVEL

CHEMISTRY

TRANSITION GUIDE

Reinforcing knowledge, skills and literacy in
Chemistry

Pearson Edexcel International Advanced Subsidiary in Chemistry (XCH11)

Pearson Edexcel International Advanced Level in Chemistry (YCH11)

First teaching September 2018

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First certification from August 2019 (International Advanced Subsidiary) and

August 2020 (International Advanced Level)



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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 (Int) GCSE. 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 International 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 (Int) GCSE and IAL and guide your students' progression.

These transition materials include:

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

The teacher version also includes answers for assessments, worksheets and exam practice questions.

The mapping of content and skills from international GCSE/GCSE to IAL should enable you to streamline your teaching and move on to the IAL 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 International GCSE/GCSE 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 International GCSE/GCSE qualifications.

Transition guide overview

Topic	Specification links	Resources
<p>Section A Atomic structure, formulae and bonding</p>	<p>IAL – Topic 2 – Atomic structure and the Periodic Table International GCSE – Core and Additional concepts</p>	<ul style="list-style-type: none"> • 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
<p>Section B Quantitative analysis and equations</p>	<p>IAL – Topic 1 – Formulae, equations and amounts of substance International GCSE – Additional and Further Additional/Extension concepts</p>	<ul style="list-style-type: none"> • 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	IAL – Topic 3 – Bonding and Structure International GCSE – Core and Additional concepts	<ul style="list-style-type: none"> • 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	Answers to Baseline assessment	
Appendix 3	Answers to worksheets	
Appendix 4	Answers to Exam practice	
Appendix 5	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: <ul style="list-style-type: none"> • 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 International GCSE 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 IAL learning.
Exam practice and Examiners' report	How to answer exam-type questions at International A 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.

NH₄Cl _____

HNO₃ _____

C₂H₄ _____

C₃H₈ _____

CO₂ _____

C₂H₅OH _____

Fe₂O₃ _____

SO₂ _____

HBr _____

NH₃ _____

(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.

SO₂ _____

KBr _____

C₂H₆ _____

Ca(OH)₂ _____

C₂H₅OH _____

NaNO₃ _____

NH₄Cl _____

FeCl₃ _____

(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. group 7 – 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.

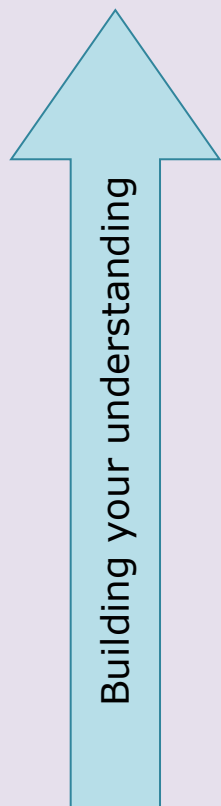
Table of resources in this section

Topics covered	Type of resource	Resource name	Brief description and notes for resource
<ul style="list-style-type: none"> 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.
<ul style="list-style-type: none"> Atomic structure Ionic compounds Electron configuration Dot-and-cross diagrams for ionic bonding Covalent compounds (simple covalent bonding) 	Teacher resource	Summary sheets	Review of (Int) GCSE concepts. Summary of key points and guide to correct use of key terms. Tips on how to answer exam questions.
<ul style="list-style-type: none"> 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.
<ul style="list-style-type: none"> Orbitals and electron configuration 	Student worksheet	Worksheet 2: Orbitals and electron configuration	Checking understanding of new IAL learning.
<ul style="list-style-type: none"> 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 IAL level. Covering main misconceptions for main topics.
<ul style="list-style-type: none"> Writing formulae Atomic structure Electron configuration Dot-and-cross diagrams 	Student questions	Exam practice	Exam questions on section covering International GCSE to IAL content. Checking how far students have progressed at the end of the section.

Teacher resources

Building knowledge

Atomic structure and formulae




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

Summary sheets

International GCSE – 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.

International GCSE/GCSE – 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*.

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

International GCSE/GCSE – 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.

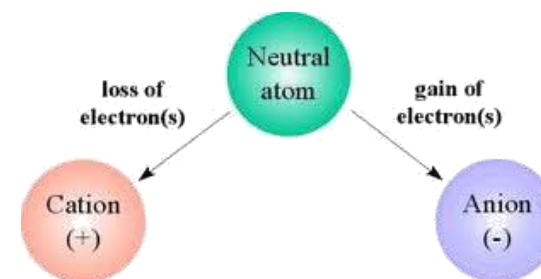
Atoms of Group 5 non-metals gain three electrons and form ions with a 3- charge, e.g. N^{3-}

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.

International GCSE/GCSE – 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					

International GCSE/GCSE – Dot-and-cross diagrams for ionic bonding

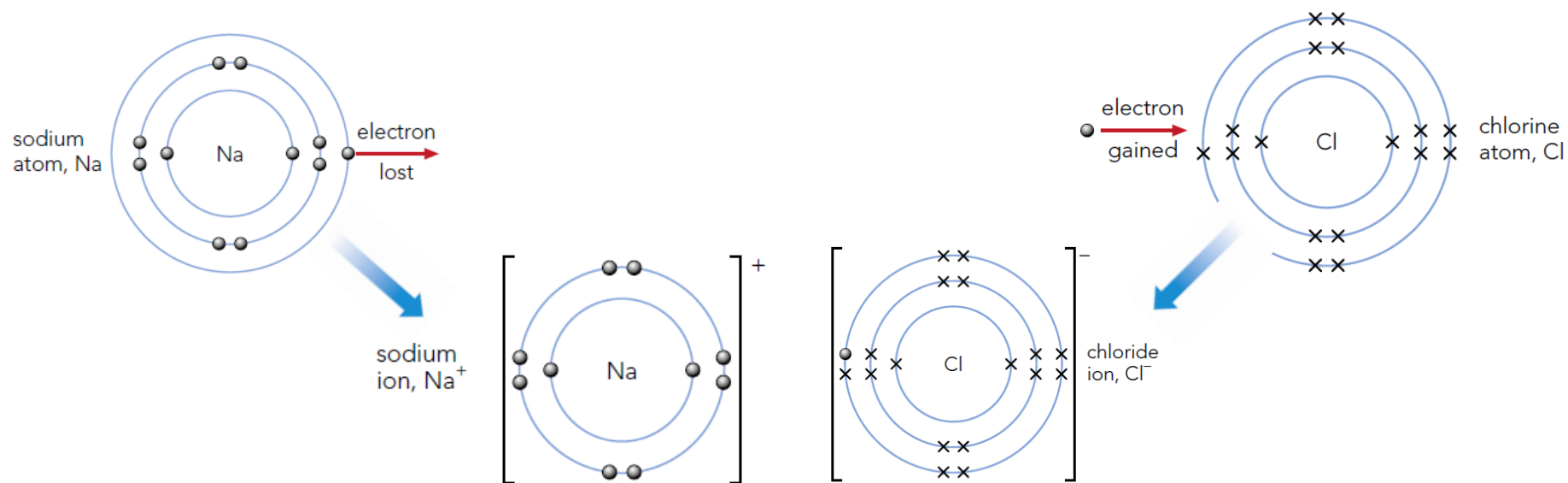
Hints and tips

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.



International GCSE/GCSE – Covalent compounds (simple covalent bonding)

A covalent bond is formed 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.

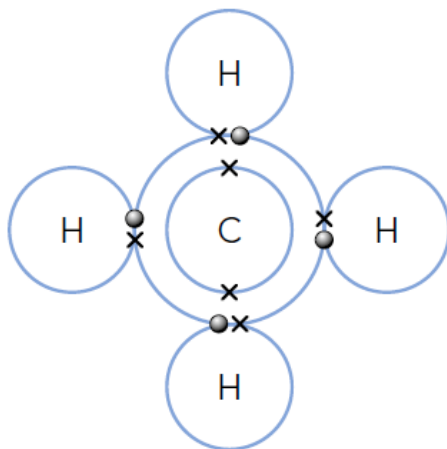
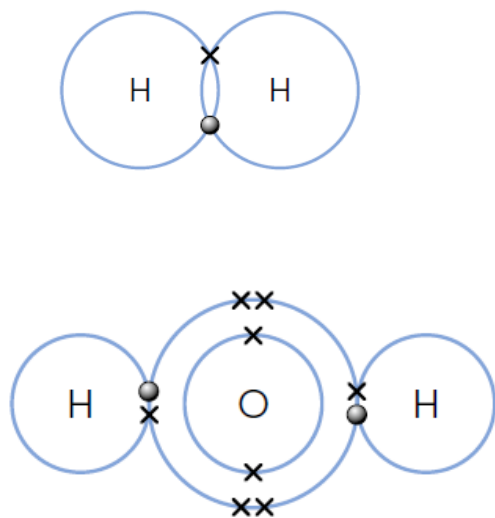
... count the final number of electrons around each atom to make sure that the outer shell is full.

Never ...

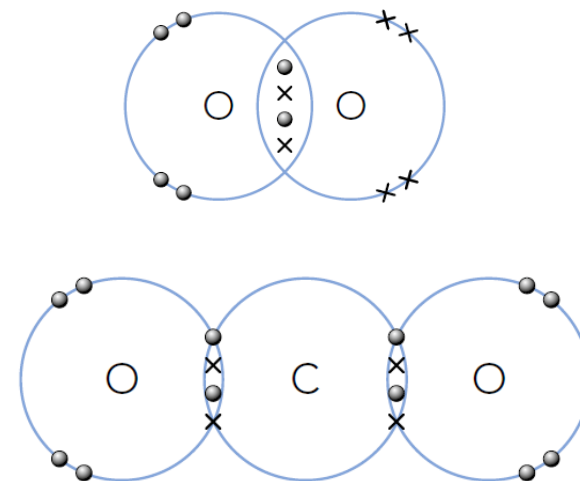
... include a charge on the atoms.

... 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.

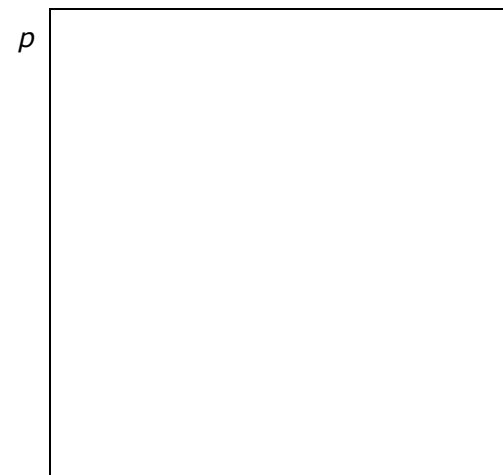
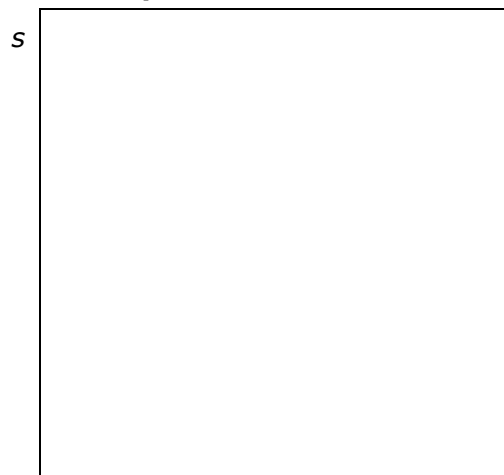
For negative ions this _____ by the number equivalent to the charge on the ion.

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.



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

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 $\frac{1}{12}$ 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.

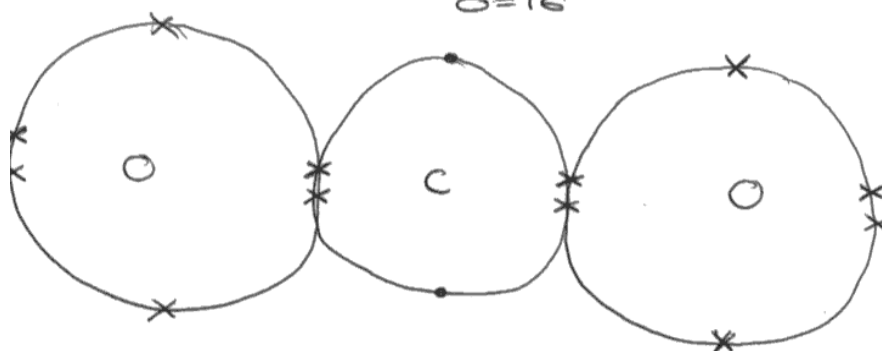
Example 3

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

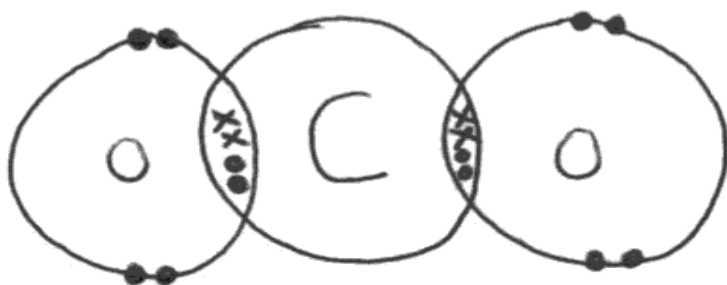
Show outer electrons only.

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 $\frac{1}{12}$ 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 different atomic structures of the same element with the same number of protons but different 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 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)

(A level question)

- 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)

(A level question)

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

(2 marks)

(GCSE question)

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

(2 marks)

(A level question)

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

(2 marks)

(GCSE question)

- 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)

(GCSE question)

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)

(A level question)

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)

(A level question)

7 The isotopes of magnesium ${}^{24}_{12}\text{Mg}$ and ${}^{25}_{12}\text{Mg}$ 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 ${}^{25}_{12}\text{Mg}^{2+}$ has more protons than ${}^{24}_{12}\text{Mg}^{2+}$.
- 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)

(A level question)

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

	1s	2s	2p		
A	$\uparrow\downarrow$	$\uparrow\downarrow$	\uparrow	\uparrow	
B	$\uparrow\downarrow$	$\uparrow\downarrow$	\uparrow	\uparrow	\uparrow
C	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	\uparrow
D	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$

(1 mark)

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

	1s	2s	2p		
A	$\uparrow\downarrow$	\uparrow	$\uparrow\downarrow$	\uparrow	\uparrow
B	$\uparrow\downarrow$	\uparrow	$\uparrow\downarrow$	$\uparrow\downarrow$	
C	$\uparrow\downarrow$	$\uparrow\downarrow$	\uparrow	\uparrow	\uparrow
D	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	\uparrow	

(1 mark)

(A level question)

- 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)

As	[Ar] 3d ¹⁰	4s	4p		
		↑↓			
Se	[Ar] 3d ¹⁰	↑↓			

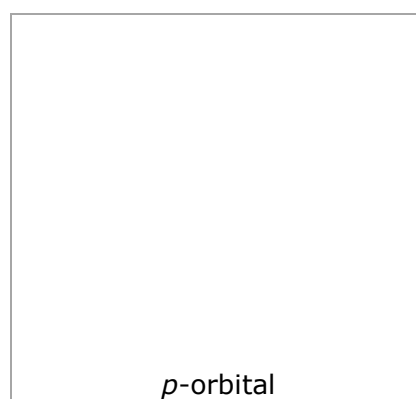
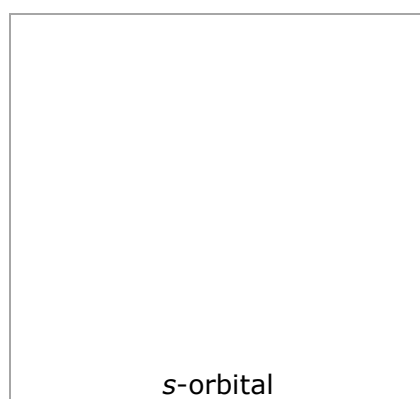
(A level question)

- 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.



(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)

(A level question)

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)

(A level question)

- 13 The radioactive isotope iodine-131, ${}_{53}^{131}\text{I}$, is formed in nuclear reactors providing nuclear power. Naturally occurring iodine contains only the isotope ${}_{53}^{127}\text{I}$.
- a** Complete the table to show the number of protons and neutrons in these two isotopes.

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

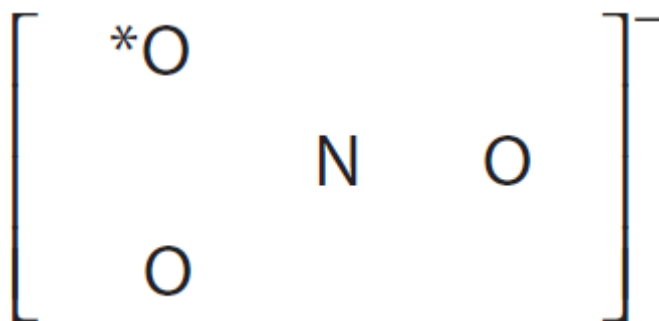
(2 marks)

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

(1 mark)

(A level question)

- 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 (x), and oxygen electrons with dots (•).
 The symbol * on the diagram represents the extra electron giving the ion its charge.



(3 marks)

(A level question)

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 A level 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 with suggestions of how answers should be laid out for clarity. 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
<ul style="list-style-type: none"> 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.
<ul style="list-style-type: none"> Writing formulae Reacting masses Percentage yield 	Teacher resource	Summary sheet: Writing formulae	Review of (Int) GCSE concepts. Summary of key points and guide to correct use of key terms. Tips on how to answer exam questions.
<ul style="list-style-type: none"> Empirical formulae Molar volumes Avogadro constant 	Teacher and student resource	Worked examples: Calculations	
<ul style="list-style-type: none"> 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 IAL learning.
<ul style="list-style-type: none"> Quantitative analysis and calculations 	Student questions	Exam practice	How to answer exam-type questions for IAL. Covering main misconceptions for main topics.

Teacher resources

Building knowledge



Quantitative chemistry

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

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

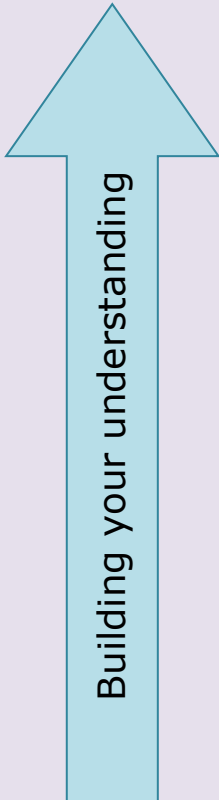
Calculate the relative atomic mass of an element given the % abundances of its isotopes.

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 → CaCl ₂	MgCO ₃ → 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 CO₂. 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.

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

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

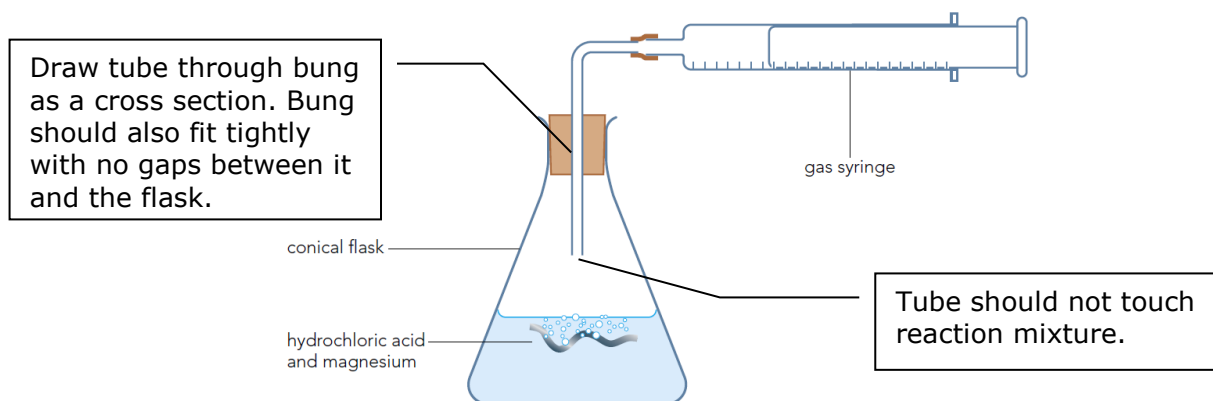
Work out empirical mass first, then use this to work out the molecular formula.

- 1 $1 \times \text{N} = 14$; $3 \times \text{H} = 3$; $1 \times \text{S} = 32$; $3 \times \text{O} = 16 \times 3 = 48$
- 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 Sulfamic 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



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

The molar volume of a gas is $24 \text{ dm}^3 \text{ mol}^{-1}$ 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 $24 \text{ dm}^3 \text{ mol}^{-1}$ at room temperature and pressure.
- 2 Number of moles of a gas = volume/molar volume.
- 3 Number of moles of H_2 = $66/24\,000 = 2.75 \times 10^{-3} \text{ mol}$.

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 $\text{H}[\text{H}_2\text{NSO}_3]$ to represent the sulfamic acid.

Interpreting the question

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

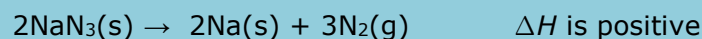
- sulfamic acid = $5.5 \times 10^{-3} \text{ mol}$.
- hydrogen molecules = $2.75 \times 10^{-3} \text{ mol}$ (answer from part **ii**).

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 $2 \text{H}[\text{H}_2\text{NSO}_3] + \text{Mg} \rightarrow \text{Mg}(\text{H}_2\text{NSO}_3)_2 + \text{H}_2$

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:



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

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 $\text{N}_2 = 50/24 = 2.08 \text{ mol}$
- 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: $2\text{NaN}_3 \rightarrow 2\text{Na} + 3\text{N}_2$
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 = $\frac{2}{3} \times 2.08 = 1.39 \text{ mol}$.
4 Convert moles to mass:
 - Molar mass of sodium azide = $23 + (14 \times 3) = 65 \text{ g mol}^{-1}$
 - Use equation: Number of moles = mass/molar mass
so mass = number of moles \times molar mass = $65 \text{ g mol}^{-1} \times 1.39 \text{ mol} = 90.4 \text{ g}$
-

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) ₂ SO 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 cm³ of hydrochloric acid (an excess) in a conical flask. When the reaction was complete, 224 cm³ of carbon dioxide had been collected over water using a 250 cm³ 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 $24\,000\text{ cm}^3\text{ mol}^{-1}$ 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.



(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 International GCSE. 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 <i>molecules</i> and <i>atoms</i> instead of <i>cations</i> or <i>ions</i> , and <i>free</i> instead of <i>delocalised</i> or <i>free-moving</i> 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 <i>free electrons</i> 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
<ul style="list-style-type: none">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.
<ul style="list-style-type: none">Ionic bonding	Teacher resource	Summary sheets	Selecting the correct vocabulary to describe bonding and properties of ionic compounds, metals and covalent compounds.
<ul style="list-style-type: none">Ionic structure	Teacher resource	Teaching ideas: Key words to describe ionic structure	Literacy activity – scaffolding resource: <ul style="list-style-type: none">how to structure long descriptive answers using the correct key wordsrelating 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.

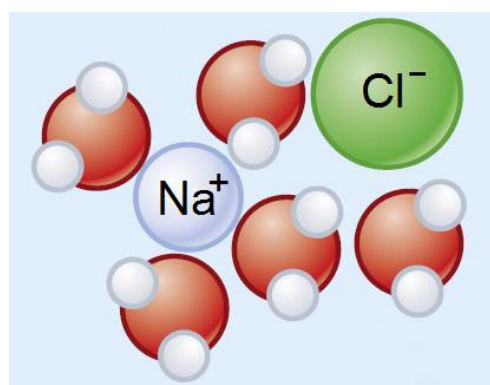
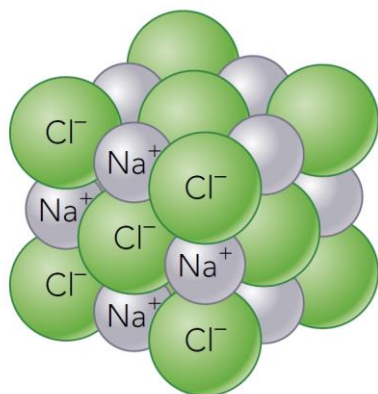
Strength of the metallic bonding increases with the number of valence electrons (outer electrons in the atoms) and with decreasing size of the cation.

Ionic bonds and ionic compounds

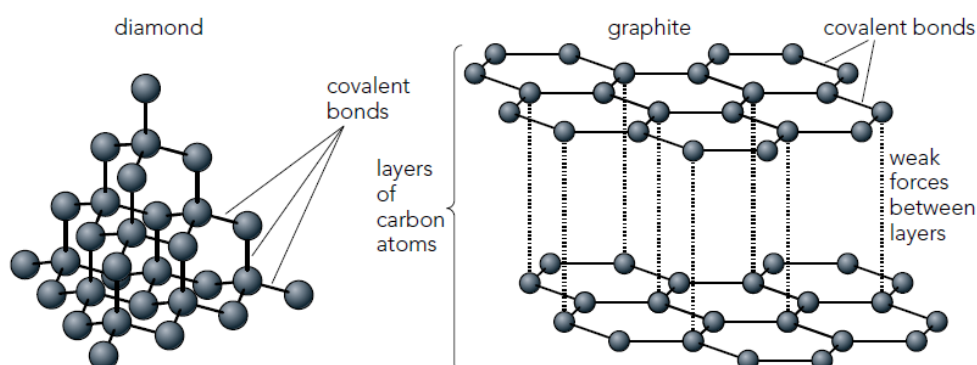
Explain why NaCl has a high melting point and only conducts electricity when molten or in solution. (5 marks)

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

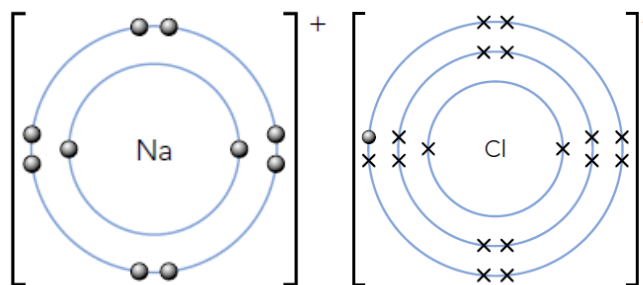
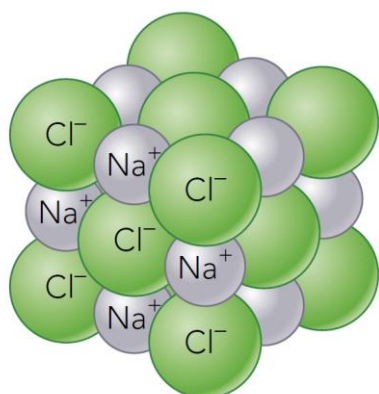


Property	Diamond	Graphite
Melting point	High – atoms held by strong covalent bonds. Many covalent bonds must be broken to melt it. Is solid at room temp.	High – atoms held by strong covalent bonds. Many covalent bonds must be broken to melt it. Is a solid at room temp.
Electrical conductivity	Poor – 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 from 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)

A level question

- 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)

A level question

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

4CH0/1C – Chemistry Paper 1
4CH0/2C – Chemistry Paper 2

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

Topic 2 – Atomic structure and the Periodic Table	International GCSE/GCSE
1. know the structure of an atom in terms of electrons, protons and neutrons	1.9 understand that atoms consist of a central nucleus, composed of protons and neutrons, surrounded by electrons, orbiting in shells
2. know the relative mass and relative charge of protons, neutrons and electrons	1.10 recall the relative mass and relative charge of a proton, neutron and electron
3. know what is meant by the terms 'atomic (proton) number' and 'mass number'	1.11 understand the terms atomic number, mass number, isotopes and relative atomic mass (A_r)
4. be able to use the atomic number and the mass number to determine the number of each type of subatomic particle in an atom or ion	1.11 understand the terms atomic number, mass number, isotopes and relative atomic mass (A_r)

Topic 2 – Atomic structure and the Periodic Table	International GCSE/GCSE
5. understand the term 'isotope'	1.11 understand the terms atomic number, mass number, isotopes and relative atomic mass (A_r)
6. understand the basic principles of a mass spectrometer and be able to analyse and interpret mass spectra to: ii calculate the relative atomic mass of an element from relative abundances of isotopes and vice versa	1.12 calculate the relative atomic mass of an element from the relative abundances of its isotopes
9. know that an orbital is a region within an atom that can hold up to two electrons with opposite spins 13. know that orbitals in sub-shells: i each take a single electron before pairing up ii pair up with two electrons of opposite spin 14. be able to predict the electronic configuration of atoms of the elements from hydrogen to krypton inclusive and their ions, using <i>s</i> , <i>p</i> , <i>d</i> notation and electron-in-boxes notation 15. understand that electronic configuration determines the chemical properties of an element	1.14 deduce the electronic configurations of the first 20 elements from their positions in the Periodic Table 1.15 deduce the number of outer electrons in a main group element from its position in the Periodic Table. 2.4 understand why elements in the same group of the Periodic Table have similar chemical properties 2.5 understand that the noble gases (Group 0) are a family of inert gases and explain their lack of reactivity in terms of their electronic configurations.

Topic 3 – Bonding and structure	International GCSE/GCSE
<p>2. be able to describe the formation of ions in terms of loss or gain of electrons</p> <p>3. be able to draw dot-and-cross diagrams to show electrons in cations and anions</p> <p>4. be able to describe ionic crystals as giant lattices of ions</p>	<p>1.28 describe the formation of ions by the gain or loss of electrons</p> <p>1.32 explain, using dot and cross diagrams, the formation of ionic compounds by electron transfer, limited to combinations of elements from Groups 1, 2, 3 and 5, 6, 7</p> <p>1.36 describe an ionic crystal as a giant three-dimensional lattice structure held together by the attraction between oppositely charged ions</p>
<p>5. know that ionic bonding is the result of strong net electrostatic attraction between ions</p>	<p>1.33 understand ionic bonding as a strong electrostatic attraction between oppositely charged ions</p>
<p>10. understand that covalent bonding is the strong electrostatic attraction between two nuclei and the shared pair of electrons between them, based on the evidence:</p> <p>i the physical properties of giant atomic structures</p> <p>ii electron density maps for simple molecules</p>	<p>1.38 describe the formation of a covalent bond by the sharing of a pair of electrons between two atoms</p> <p>1.39 understand covalent bonding as a strong attraction between the bonding pair of electrons and the nuclei of the atoms involved in the bond</p>

Topic 3 – Bonding and structure	International GCSE/GCSE
<p>11. be able to draw dot-and-cross diagrams to show electrons in covalent substances, including:</p> <ul style="list-style-type: none"> i molecules with single, double and triple bonds ii species with dative covalent (coordinate) bonds, including Al_2Cl_6 and the ammonium ion <p>12. be able to describe the different structures formed by giant lattices of carbon atoms, including graphite, diamond and graphene, and discuss the applications of each</p>	<p>1.40 explain, using dot and cross diagrams, the formation of covalent compounds by electron sharing for the following substances:</p> <ul style="list-style-type: none"> a) hydrogen b) chlorine c) hydrogen chloride d) water e) methane f) ammonia g) oxygen h) nitrogen i) carbon dioxide j) ethane k) ethane <p>1.44 draw diagrams representing the positions of the atoms in diamond and graphite</p> <p>1.45 explain how the uses of diamond and graphite depend on their structures, limited to graphite as a lubricant and diamond in cutting.</p>
<p>20. understand that metals consist of giant lattices of metal ions in a sea of delocalised electrons</p> <p>21. know that metallic bonding is the strong electrostatic attraction between metal ions and the delocalised electrons</p> <p>22. be able to use the models in 3.20 and 3.21 to interpret simple properties of metals, including electrical conductivity and high melting temperature</p>	<p>1.46 understand that a metal can be described as a giant structure of positive ions surrounded by a sea of delocalised electrons</p> <p>1.47 explain the electrical conductivity and malleability of a metal in terms of its structure and bonding.</p>

Topic 1 – Formulae, equations and amounts of substance	International GCSE/GCSE
<p>6. be able to define the terms <i>relative isotopic mass</i> and <i>relative atomic mass</i>, based on the ^{12}C scale</p> <p>7. understand the terms <i>relative molecular mass</i> and <i>relative formula mass</i>, including calculating these values from relative atomic masses</p> <p>Definitions of these terms will not be expected</p> <p>The term <i>relative formula mass</i> should be used for compounds with giant structures</p>	<p>1.11 understand the terms atomic number, mass number, isotopes and relative atomic mass (A_r)</p> <p>1.12 calculate the relative atomic mass of an element from the relative abundances of its isotopes</p>
<p>2. know that the mole (mol) is the unit for the amount of a substance and be able to perform calculations using the Avogadro constant L ($6.02 \times 10^{23} \text{ mol}^{-1}$)</p>	<p>1.17 understand the use of the term mole to represent the amount of substance</p> <p>1.18 understand the term mole as the Avogadro number of particles (atoms, molecules, formulae, ions or electrons) in a substance</p>

Topic 1 – Formulae, equations and amounts of substance	International GCSE/GCSE
<p>4. understand the terms:</p> <ul style="list-style-type: none"> i 'relative atomic mass' based on the ^{12}C scale ii 'relative molecular mass' and 'relative formula mass', including calculating these values from relative atomic masses <i>The term 'relative formula mass' should be used for compounds with giant structures</i> iii 'molar mass' as the mass per mole of a substance in g mol^{-1} 	<p>1.19 carry out mole calculations using relative atomic mass (A_r) and relative formula mass (M_r)</p>
<p>6. be able to use experimental data to calculate empirical and molecular formulae</p>	<p>1.23 understand how the formulae of simple compounds can be obtained experimentally, including metal oxides, water and salts containing water of crystallisation</p> <p>1.24 calculate empirical and molecular formulae from experimental data</p>
<p>3. write balanced full and ionic equations, including state symbols, for chemical reactions</p>	<p>1.21 write word equations and balanced chemical equations to represent the reactions studied in this specification</p> <p>1.22 use the state symbols (s), (l), (g) and (aq) in chemical equations to represent solids, liquids, gases and aqueous solutions respectively</p>
<p>7. be able to use chemical equations to calculate reacting masses and vice versa, using the concepts of amount of substance and molar mass</p>	<p>1.25 calculate reacting masses using experimental data and chemical equations</p> <p>1.27 carry out mole calculations using volumes and molar concentrations.</p>

Topic 1 – Formulae, equations and amounts of substance	International GCSE/GCSE
<p>8. be able to use chemical equations to calculate volumes of gases and vice versa, using:</p> <ul style="list-style-type: none"> i the concepts of amount of substance ii the molar volume of gases iii the expression $pV = nRT$ for gases and volatile liquids <p>11. Core Practical 1</p>	<p>1.20 understand the term molar volume of a gas and use its values (24 dm³ and 24,000 cm³) at room temperature and pressure (rtp) in calculations.</p>
<p>9. be able to calculate percentage yields and percentage atom economies (by mass) in laboratory and industrial processes, using chemical equations and experimental results</p> <p>Atom economy = $\frac{\text{molar mass of the desired product}}{\text{sum of the molar masses of all products}} \times 100\%$</p>	<p>1.26 calculate percentage yield</p>
<p>5. calculate the concentration of a solution in mol dm⁻³ and g dm⁻³</p>	<p>1.27 carry out mole calculations using volumes and molar concentrations.</p>

Topic 1 – Formulae, equations and amounts of substance	International GCSE/GCSE
<p>12. be able to relate ionic and full equations, with state symbols, to observations from simple test-tube experiments, to include:</p> <ul style="list-style-type: none"> i displacement reactions ii typical reactions of acids iii precipitation reactions 	<p>2.38 describe tests for the anions:</p> <ul style="list-style-type: none"> i Cl^-, Br^- and I^-, using dilute nitric acid and silver nitrate solution ii SO_4^{2-}, using dilute hydrochloric acid and barium chloride solution iii CO_3^{2-}, using dilute hydrochloric acid and identifying the carbon dioxide evolved <p>4.6 understand the general rules for predicting the solubility of salts in water:</p> <ul style="list-style-type: none"> i all common sodium, potassium and ammonium salts are soluble ii all nitrates are soluble iii common chlorides are soluble, except silver chloride iv common sulfates are soluble, except those of barium and calcium v common carbonates are insoluble, except those of sodium, potassium and ammonium

Appendix 2: Answers to Baseline assessment

- 1** Copper(II) sulfate – CuSO_4
 Sodium hydroxide – NaOH
 Strontium nitrate – $\text{Sr}(\text{NO}_3)_2$
 Sodium carbonate – Na_2CO_3
 Lithium hydrogencarbonate – LiHCO_3
 Potassium nitrate – KNO_3
 Calcium hydroxide – $\text{Ca}(\text{OH})_2$
 Aluminium fluoride – AlF_3

0–1 correct score = 0 marks
2–3 correct = 1 mark
4–5 correct = 2 marks
6–7 correct = 3 marks
All correct = 4 marks

- 2** NH_4Cl – Ammonium chloride
 C_2H_4 – Ethene
 CO_2 – Carbon dioxide
 Fe_2O_3 – Iron(III) oxide
 HBr – hydrogen bromide
 HNO_3 – Nitric acid
 C_3H_8 – Propane
 $\text{C}_2\text{H}_5\text{OH}$ – Ethanol
 SO_2 – Sulfur dioxide
 NH_3 – Ammonia

0–1 correct score = 0 marks
2–3 correct = 1 mark
4–5 correct = 2 marks
6–7 correct = 3 marks
8–9 = 4 marks
All correct = 5 marks

- 3** 1 mark for each correct row.

Particle	Where it is found	Charge	Mass
Neutron	nucleus	0	1
Proton	nucleus	+1	1
Electron	Electron shells/around the nucleus	-1	0

- 4
- | | |
|----------------------------------|--|
| SO ₂ | $32.1 + (2 \times 16.0) = 64.1$ |
| C ₂ H ₆ | $(2 \times 12.0) + (6 \times 1.0) = 30.0$ |
| C ₂ H ₅ OH | $(2 \times 12.0) + (5 \times 1.0) + 16 + 1 = 46.0$ |
| NH ₄ Cl | $14.0 + (4 \times 1.0) + 35.5 = 53.5$ |
| KBr | $39.1 + 79.9 = 119$ |
| Ca(OH) ₂ | $40.1 + (2 \times (1.0 + 16.0)) = 74.1$ |
| NaNO ₃ | $23.0 + 14.0 + (3 \times 16.0) = 85.0$ |
| FeCl ₃ | $55.8 + (35.5 \times 3) = 162.3$ |

- 5
- a The sum of the number of protons and the number of neutrons in the nucleus of an atom. (1 mark)
- b The weighted mean mass of an atom of the element compared to 1/12 th of the mass the mass of an atom of carbon-12 (1 mark), which has a mass of 12. (1 mark).
- c Atoms of the same element that have different masses. (1 mark).
- 6 1 mark for each word equation. 2 marks for each symbol equation.

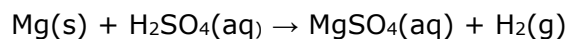
Calcium carbonate and hydrochloric acid

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



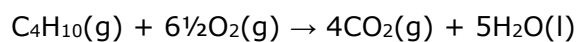
Magnesium and sulfuric acid

magnesium and sulphuric acid → magnesium sulphate + hydrogen

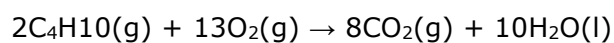


Complete combustion of butane

butane + oxygen → carbon dioxide + water



OR



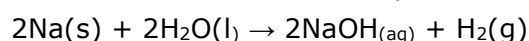
Thermal decomposition of calcium carbonate

calcium carbonate → calcium oxide + carbon dioxide



Sodium and water

sodium + water → sodium hydroxide + hydrogen gas



7 Ionic bonding: Electrostatic attraction between oppositely charged ions. (1 mark)

Covalent bonding: Electrostatic attraction between a bonding pair of electrons and the nuclei of the two bonded atoms (1 mark)

Metallic bonding: Electrostatic attraction between the nuclei of positive ions (cations) and delocalised electrons (1 mark).

8 1 mark for each 3 correct answers.

Substance	Formula	Type of bonding	Type of structure
Hydrogen sulfide	H₂S	covalent	(simple) molecular
Graphite	C	covalent	giant
Silicon dioxide	SiO₂	covalent	giant
Methane	CH₄	covalent	(simple) molecular
Calcium	Ca	metallic	giant
Magnesium chloride	MgCl₂	ionic	giant

9 1 mark for each point.

- Gas molecules are adsorbed onto the layers in graphite
 - which allow the layers to slide past each other.
 - The delocalised electrons between the layers
 - can flow under the influence of a potential difference (applied parallel to the layers).
-

Appendix 3: Answers to worksheets

Section A

Worksheet 1 answers

Structure of an atom

The nucleus contains protons and neutrons.

The electrons are found in the electron shells or energy levels.

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 the atomic number.

In a neutral atom the number of electrons is equal to the number of protons.

To work out the number of neutrons we subtract the atomic number from the mass number.

Structure of an ion

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

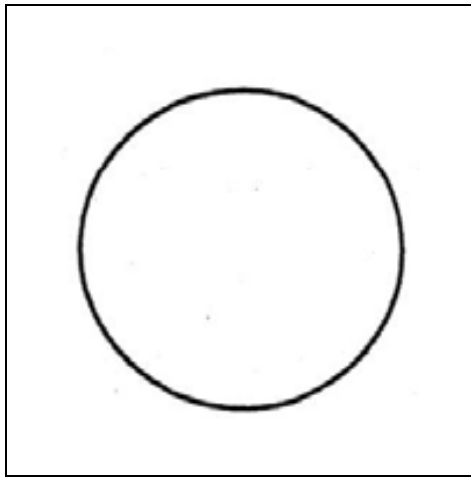
For positive ions this decreases by the number equivalent to the charge on the ion.

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

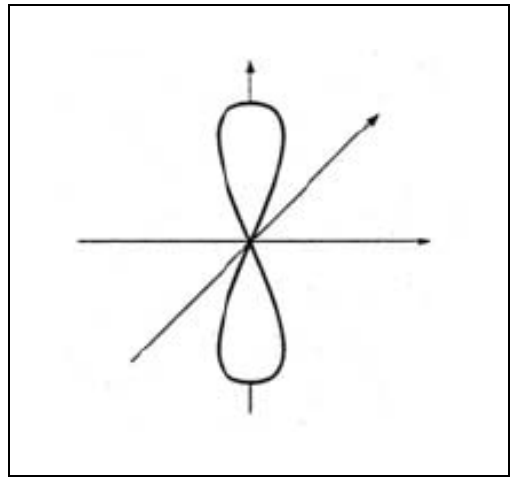
Worksheet 2 answers

Quantum shell	Maximum number of electrons	Types of orbitals	Total number of orbitals	Electron configuration
$n = 1$	2	s	1	$1s^2$
$n = 2$	8	s, p	4	$2s^2 2p^6$
$n = 3$	18	s, p, d	9	$3s^2 3p^6 3d^{10}$
$n = 4$	32	s, p, d, f	16	$4s^2 4p^6 4d^{10} 4f^{14}$

s



ρ



		Electron configuration		Electrons in boxes														
	Z	2.8.8	s,p,d	1s	2s	2p			3s	3p			3d					4s
Na	11	2.8.1	$1s^2 2s^2 sp^6 3s^1$	↑↓	↑↓	↑↓	↑↓	↑↓	↑									
Be	4	2.2	$1s^2 2s^2$	↑↓	↑↓													
Be²⁺	4	2	$1s^2$	↑↓														
P	15	2.8.5	$1s^2 2s^2 2p^6 3s^2 3p^3$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑						
Cr	24	2.8.13.1	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑	↑	
Cu	29	2.8.18.1	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	
Fe	26	2.8.14.2	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑	
Al	13	2.8.3	$1s^2 2s^2 2p^6 3s^2 3p^1$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑								
Al³⁺	13	2.8	$1s^2 2s^2 2p^6$	↑↓	↑↓	↑↓	↑↓	↑↓										
Sc	21	2.8.9.2	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑				↑↓	
Cl	17	2.8.7	$1s^2 2s^2 2p^6 3s^2 3p^5$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑						
Cl⁻	17	2.8.8	$1s^2 2s^2 2p^6 3s^2 3p^6$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓						

Section B

Worksheet 1 answers

Copper(II) sulfate	CuSO₄
Nitric acid	HNO₃
Copper(II) nitrate	Cu(NO₃)₂
Sulfuric acid	H₂SO₄
Sodium carbonate	Na₂CO₃
Aluminium sulfate	Al₂(SO₄)₃
Ammonium nitrate	NH₄NO₃
Nitrogen dioxide	NO₂
Sulfur dioxide	SO₂
Ammonia	NH₃
Ammonium sulfate	(NH₄)₂SO₄
Potassium hydroxide	KOH
Calcium hydroxide	Ca(OH)₂

Worksheet 2 answers

Substance	Formula	Cation (exact number)	Anion (exact number)
Sodium bromide	NaBr	Na⁺	Br⁻
Potassium iodide	KI	K⁺	I⁻
Silver nitrate	AgNO₃	Ag⁺	NO₃⁻
Copper(II) sulfate	CuSO₄	Cu²⁺	SO₄²⁻
Sodium hydrogencarbonate	NaHCO ₃	Na⁺	HCO₃⁻
Magnesium carbonate	MgCO₃	Mg²⁺	CO₃²⁻
Lithium carbonate	Li₂CO₃	2Li⁺	CO₃²⁻
Calcium hydrogensulfate	Ca(HSO ₄) ₂	Ca²⁺	HSO₄⁻
Aluminium nitrate	Al(NO₃)₃	Al³⁺	3NO₃⁻
Calcium phosphate	Ca₃(PO₄)₂	3Ca²⁺	2PO₄³⁻
Potassium hydride	KH	K⁺	H⁻
Sodium ethanoate	CH₃COONa	Na⁺	CH₃COO⁻
Potassium manganate(VII)	KMnO ₄	K⁺	MnO₄⁻
Potassium dichromate(VI)	K₂Cr₂O₇	2K⁺	Cr₂O₇²⁻
Zinc chloride	ZnCl₂	Zn²⁺	2Cl⁻

Substance	Formula	Cation (exact number)	Anion (exact number)
Strontium nitrate	Sr(NO₃)₂	Sr²⁺	2NO₃⁻
Sodium chromate(VI)	NaCrO₄	Na⁺	CrO₄²⁻
Calcium fluoride	CaF₂	Ca²⁺	2F⁻
Potassium sulfide	K₂S	2K⁺	S²⁻
Magnesium nitride	Mg₃N₂	3Mg²⁺	2N³⁻
Lithium hydrogensulfate	Li(HSO₄)	Li⁺	HSO₄⁻
Ammonium sulfate	(NH ₄) ₂ SO ₄	2NH₄⁺	SO₄²⁻

Worksheet 3 answers

- $$\text{Mg(s)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{MgSO}_4\text{(aq)} + \text{H}_2\text{(g)}$$

$$\text{Mg(aq)} + 2\text{H}^+\text{(aq)} \rightarrow \text{Mg}^{2+}\text{(aq)} + \text{H}_2\text{(g)}$$
 - $$\text{CaCO}_3\text{(s)} + 2\text{HNO}_3\text{(aq)} \rightarrow \text{Ca(NO}_3)_2\text{(aq)} + \text{H}_2\text{O(l)}$$

$$\text{CaCO}_3\text{(s)} + 2\text{H}^+\text{(aq)} \rightarrow \text{Ca}^{2+}\text{(aq)} + \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)}$$
 - $$\text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$$

$$\text{H}^+\text{(aq)} + \text{OH}^-\text{(aq)} \rightarrow \text{H}_2\text{O(l)}$$
 - $$\text{BaCl}_2\text{(aq)} + \text{Na}_2\text{SO}_4\text{(aq)} \rightarrow \text{BaSO}_4\text{(s)} + 2\text{NaCl(aq)}$$

$$\text{Ba}^{2+}\text{(aq)} + \text{SO}_4^{2-}\text{(aq)} \rightarrow \text{BaSO}_4\text{(s)}$$
 - $$2\text{NaOH(aq)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{Na}_2\text{SO}_4\text{(aq)} + \text{H}_2\text{O(l)}$$

$$\text{H}^+\text{(aq)} + \text{OH}^-\text{(aq)} \rightarrow \text{H}_2\text{O(l)}$$
 - $$2\text{AgNO}_3\text{(aq)} + \text{MgCl}_2\text{(aq)} \rightarrow 2\text{AgCl(s)} + \text{Mg(NO}_3)_2\text{(aq)}$$

$$\text{Ag}^+\text{(aq)} + \text{Cl}^-\text{(aq)} \rightarrow \text{AgCl(s)}$$
 - $$\text{MgO(s)} + 2\text{HNO}_3\text{(aq)} \rightarrow \text{Mg(NO}_3)_2\text{(aq)} + \text{H}_2\text{O(l)}$$

$$\text{MgO(s)} + 2\text{H}^+\text{(aq)} \rightarrow \text{Mg}^{2+}\text{(aq)} + \text{H}_2\text{O(l)}$$
 - $$\text{CuSO}_4\text{(aq)} + 2\text{NaOH(aq)} \rightarrow \text{Cu(OH)}_2\text{(s)} + \text{Na}_2\text{SO}_4\text{(aq)}$$

$$\text{Cu}^{2+}\text{(aq)} + 2\text{OH}^-\text{(aq)} \rightarrow \text{Cu(OH)}_2\text{(s)}$$
 - $$\text{Pb(NO}_3)_2\text{(aq)} + 2\text{KI(aq)} \rightarrow \text{PbI}_2\text{(s)} + 2\text{KNO}_3\text{(aq)}$$

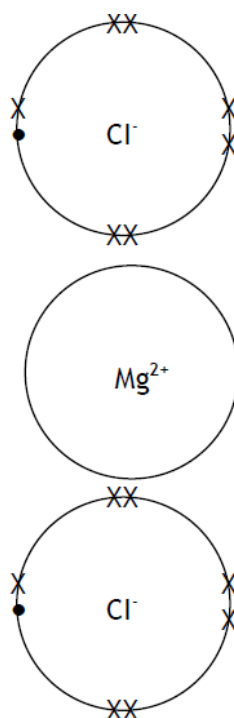
$$\text{Pb}^{2+}\text{(aq)} + 2\text{I}^-\text{(aq)} \rightarrow \text{PbI}_2\text{(s)}$$
 - $$\text{Fe}_2\text{(SO}_4)_3 + 6\text{NaOH(aq)} \rightarrow 3\text{Fe(OH)}_3\text{(s)} + 3\text{Na}_2\text{SO}_4\text{(aq)}$$

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

Appendix 4: Answers to Exam practice

Section A

- 1 The (weighted) mean mass of an atom of the element (1 mark) compared to 1/12 th of the mass the mass of an atom of carbon-12 (1 mark) (which has a mass of 12).
- 2 a Correct number of outer electrons (ignore whether dots and/or crosses) drawn and also ratio of magnesium : chloride ions is 1:2 (1 mark).



Correct formulae and charges of the ions shown somewhere (1 mark).

Note: Diagram for Mg²⁺ showing the outermost shell with 8e⁻ (dots and/or crosses) and/or Cl⁻ shown with a 2 in front or 2 as a subscript would also score both marks.

- b 4 shared pairs of electrons around the carbon labelled C (1 mark).
All outer electrons, including lone pairs, are correctly shown on each of the four chlorine atoms labelled Cl (1 mark).
Allow versions without circles.
- c Diagram showing:
- any shared pair of electrons between a carbon and oxygen atom in CO₂ molecule (1 mark)
 - rest of molecule correct (1 mark).
- Must have O C O arrangement.
If any atom labelled must be correct.

- 3 a** The relative isotopic mass is the mass of an atom of an isotope of the element (1 mark) compared to $\frac{1}{12}$ th the mass the mass of an atom of carbon-12 (1 mark) (which has a mass of 12).
- b** Atoms of the same element / atoms with the same atomic number / atoms with the same number of protons (1 mark) with different masses / with different mass numbers / with different numbers of neutrons (1 mark)
- c i** The (weighted) mean mass of an atom of the element (1 mark) compared to $\frac{1}{12}$ th of the mass the mass of an atom of carbon-12 (1 mark) (which has a mass of 12).
- ii** $(19.7 \times 10) + (80.3 \times 11)$ (1 mark) $\div 100$ (1 mark) = 10.8 (1 mark)
OR
 (0.197×10) (1 mark) + (0.803×11) (1 mark) = 10.8 (1 mark)

4 B

5 D

$$\frac{1}{12}$$

6 B

7 C

8 D

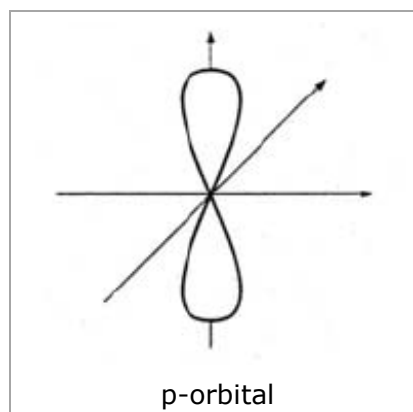
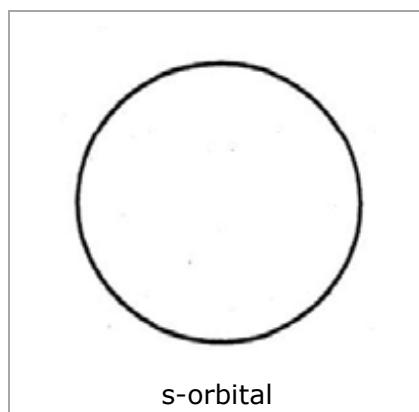
9 D

10 C

- 11 a** One mark for each row. Allow two arrows for Se in any 4p box.

As	[Ar] 3d ¹⁰	4s	4p		
		↑↓	↑	↑	↑
Se	[Ar] 3d ¹⁰	↑↓	↑↓	↑	↑

- b i** A region (around the nucleus) where there is a high probability of finding an electron.
Or
A region (around the nucleus) that can hold up to two electrons (with opposite spin).
- ii** 1 mark for each diagram. Allow 1 mark for correct diagrams in wrong boxes.



c 11/eleven (1 mark)

d 18/eighteen (1 mark)

12 a $(194 \times 32.8) + (195 \times 30.6) + (53.0 \times 25.4) + (198 \times 11.2) \div 100$ (1 mark)
 $= 195.262$
 $= 195.3$ (1 dp) (1 mark)

b d(-block) (1 mark)

13 a 1 mark for each correct row.

Isotope	$\begin{matrix} 131 \\ 53 \end{matrix}$	$\begin{matrix} 127 \\ 53 \end{matrix}$
Number of protons	53	53
Number of neutrons	78	74

(2 marks)

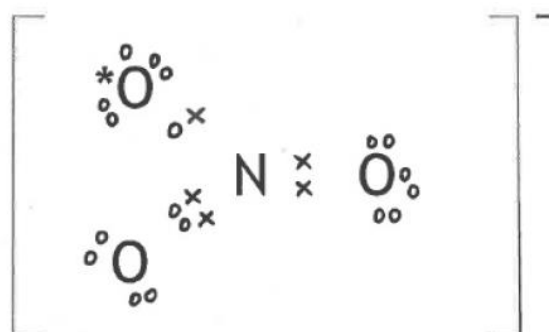
b Xenon/Xe (1 mark)

14 Double bond between N and one oxygen atom (1 mark)

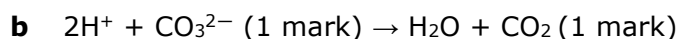
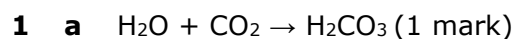
Single bond between N and O* (1 mark)

Dative single bond between N and one O atom (1 mark)

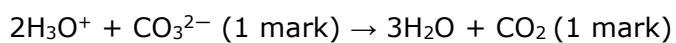
Maximum of 2 marks if any lone pair electrons are missing from any of the three oxygen atoms.



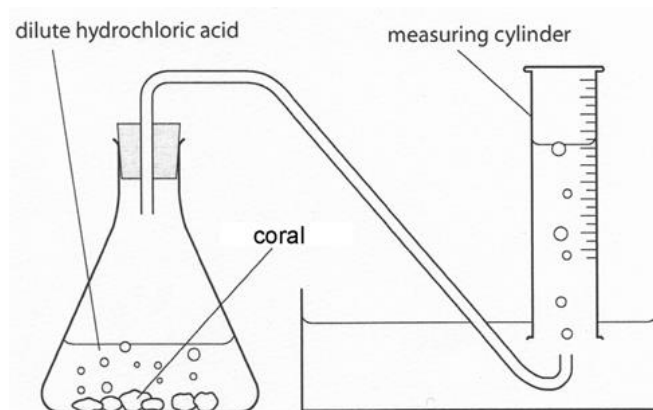
Section B



OR



- 2 a Conical flask and a delivery tube leaving the conical flask (1 mark). Inverted measuring cylinder with collection over water shown and cylinder above mouth of delivery tube (1 mark).



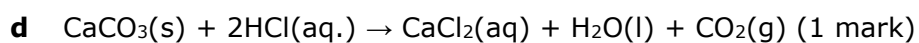
- b Any method which is likely to bring the reactants into contact after the apparatus is sealed (1 mark).

Reject: Method suggesting mixing the reactants and then putting bung in flask very quickly.

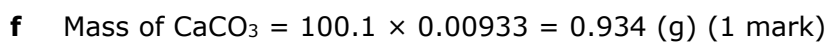
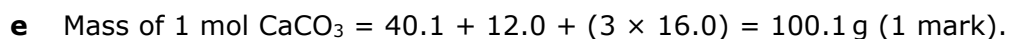
- c $(224 \div 24000 =) 0.00933 = 9.33 \times 10^{-3}$ (mol) (1 mark).

Ignore SF except 1 SF. **Ignore** any incorrect units.

Reject '0.009' as answer.

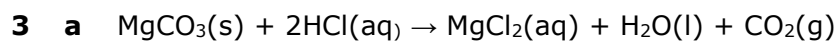


All four state symbols must be correct for this mark



Percentage of CaCO_3 in the coral = $100 \times 0.9334 / 1.13 = 82.6\%$ (1 mark)

g Different samples of coral have different amounts of CaCO_3 / different proportions of CaCO_3 (1 mark).

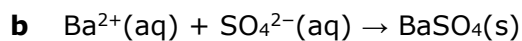


OR



All formulae and balancing (1 mark).

Mark state symbols independently (1 mark)



Species (1 mark).

State symbols (1 mark).

Section C

- 1** The oxide ion has a greater (negative) charge / greater charge density than the chloride ion (1 mark)
so the force of attraction between ions is stronger in MgO (than in MgCl₂) / stronger ionic bonding in MgO (than in MgCl₂) (1 mark)
therefore more energy is required to overcome the forces of attraction in MgO (than in MgCl₂) / more energy is required to break the (ionic) bonds in MgO (than in MgCl₂). (1 mark)
- 2 a** Silicon's (outer) electrons are fixed (in covalent bonds) / silicon's (outer) electrons are in fixed positions (in covalent bonds) / silicon's (outer) electrons are involved in bonding (1 mark).
therefore silicon's electrons are not free to move / there are no mobile electrons in silicon / silicon has no delocalized electrons / silicon's electrons cannot flow (1 mark).
- b** The covalent bonds are strong (1 mark)
therefore a lot of energy is required to break the bonds (1 mark).
- 3 a** Sodium ions are larger (than magnesium ions). (1 mark)
Sodium has fewer delocalised electrons (than magnesium). (1 mark)
Attraction between the (nuclei of) positive ions and delocalised electrons is weaker in sodium (than magnesium). (1 mark)
Allow reverse arguments in each case.
- b.** In silicon strong covalent bonds have to be broken in silicon (1 mark)
In white phosphorus weak intermolecular forces / weak London forces / weak dispersion forces / weak instantaneous dipole-induced dipole forces have to be overcome. (1 mark)
More energy needed to break the covalent bonds in silicon (than overcome the intermolecular forces in white phosphorus). (1 mark)
- c** Argon is consists of monatomic molecules / argon is composed of single atoms.
- d** Magnesium has delocalised electrons that are free to flow (under the influence of a potential difference) (1 mark)
Sulfur's (outer) electrons are fixed in covalent bonds / sulfur's (outer) electrons are involved in bonding / sulfur's (outer) electrons are not free to flow / there are no delocalised electrons in sulfur / there are no mobile electrons in sulfur (1 mark).
-

Appendix 5: 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				
O²⁻				
H⁺				
Kr				
Al³⁺				

(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

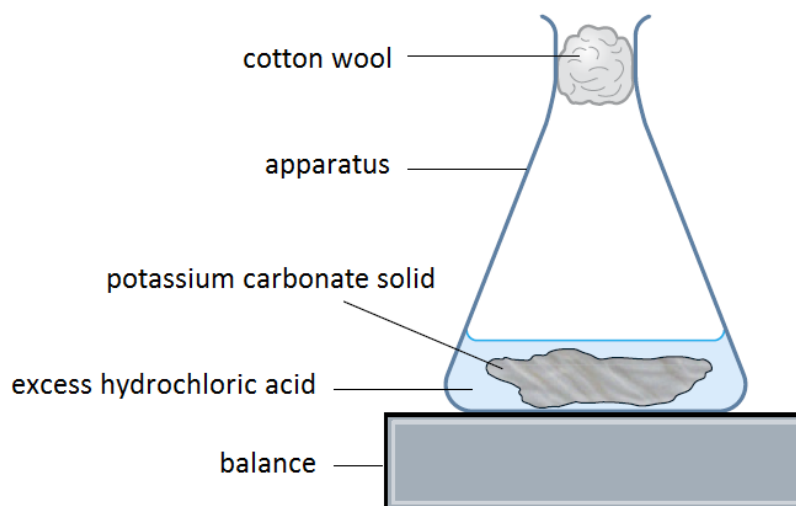
(18 marks)

Section B: baseline assessment extra questions

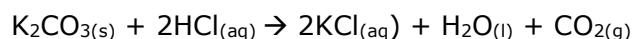
- 1 Magnesium has three isotopes. The mass spectrum of magnesium shows peaks at m/z 24 (78.60%), 25 (10.11%) and 26 (11.29%). Calculate the relative atomic mass of magnesium to 4 significant figures.

(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:



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)
