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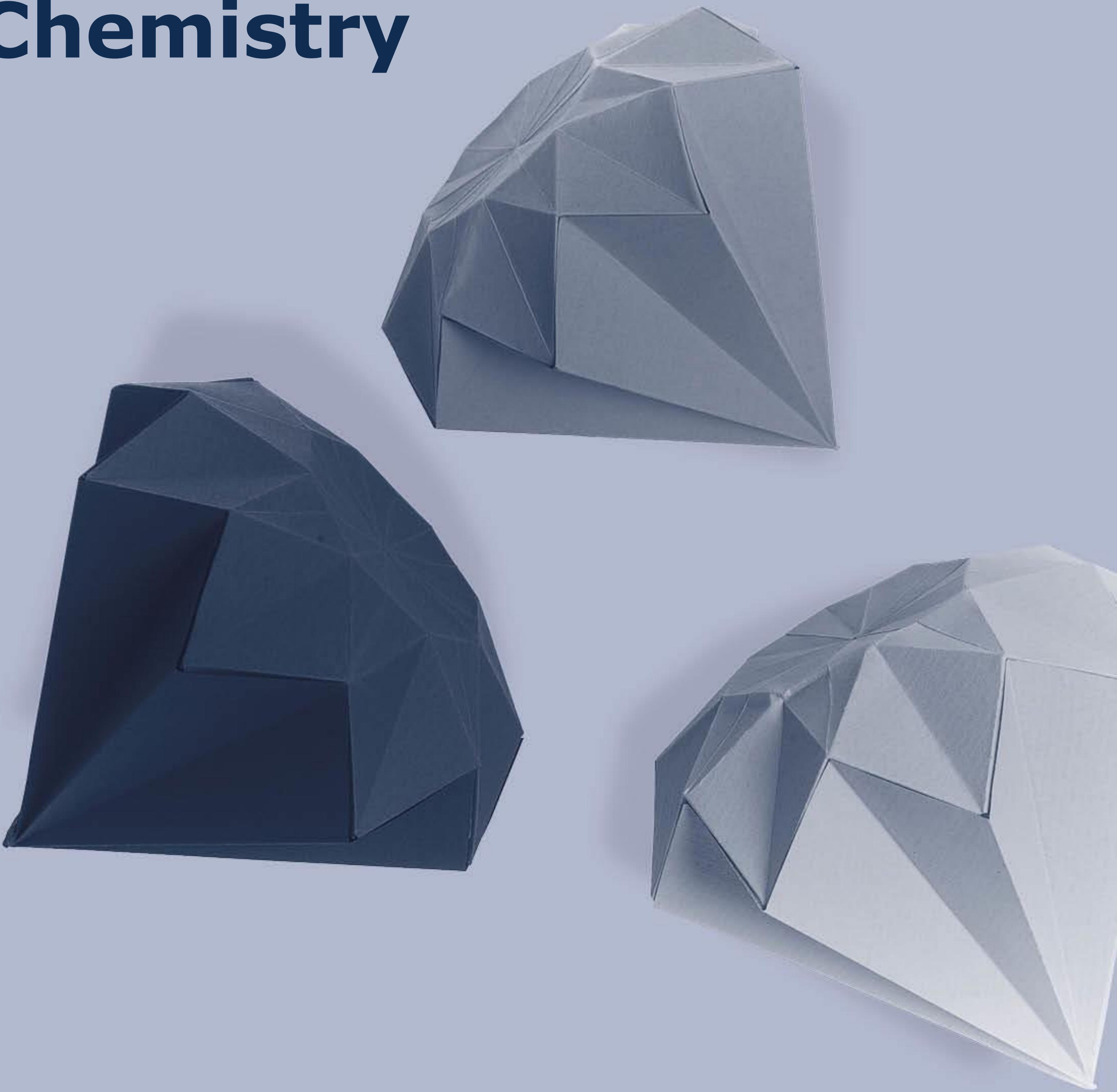
AS and A Level Chemistry



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 teacher version also includes answers for assessments, worksheets and exam 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 Year 11 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 Topic 0: Formulae, equations and hazards Topic 1: Key Concepts in Chemistry SC/CC 3 &4 – Atomic structure, the periodic table, ionic bonding, covalent bonding	• 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 Topic 1: Key Concepts in Chemistry SC9/CC9 – Calculations with masses Plus Separate Chemistry Topic 5, SC14 Quantitative Analysis	• 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 Topic 1: Key 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	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: • 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.

Student's Checklist for Being A Level Chemistry-ready

You are expected to know/understand the following:

Electron configuration for the first 20 elements

Naming compounds from formulae and vice versa

Bonding

Three main types – formation and properties

- ionic
- covalent
- and metallic

Dot and cross diagrams for covalent molecules and ionic compounds to include:

- sodium chloride, calcium oxide, calcium fluoride, aluminium oxide
- chlorine, oxygen, nitrogen, ammonia, carbon dioxide, methane, ethane, ethanol, sulfur dioxide, water

Writing formulae for all of the above plus compounds with:

- carbonate
- nitrate
- hydroxide
- hydrogen carbonate
- sulfate.

Balancing equations for

- neutralisation
- metals with acids
- alkali metals with water
- redox (displacement of halogens and metals), thermal decomposition

Calculations

- relative atomic mass, relative formula mass and empirical formulae
- Percentage yield and atom economy
- Reacting masses and limiting reagent

Energetics

- difference between exothermic and endothermic
- graphs associated with these
- energies in bond making and bond breaking

Organic Chemistry

- differences between alkanes and alkenes
- naming and reactions of alkanes and alkenes
- fractional distillation
- cracking
- characteristics of good fuels
- balancing combustion equations

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 GCSE Topic 1: Key Concepts. 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 learned at KS4. With the GCSE 2016 specifications, this is an area that is assessed in both papers in the exams. The move to linear exams means that the gap in retention of knowledge from GCSE to GCE should be narrower. 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. It is expected that students will be more 'A level-ready' than in previous years.

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

Summary sheets

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*.

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

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+}

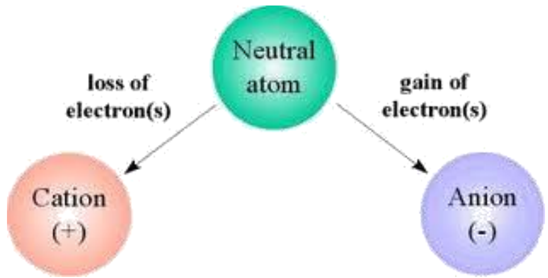
Atoms of non-metallic elements in Groups 5, 6 and 7 can form negative ions when they take part in reactions since they are able to gain electrons.

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					

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.

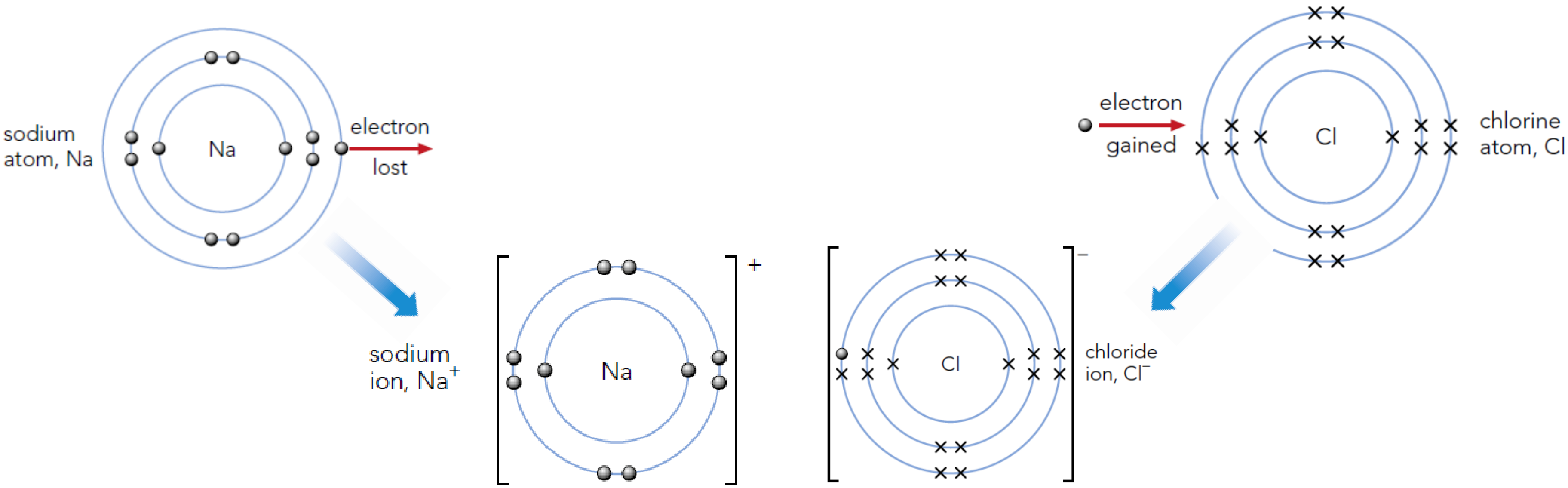
Never …

… 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)

Distinguish between:

'How a covalent bond is formed': A covalent bond is formed when a pair of electrons is shared between two atoms.

'What is a covalent bond?': Electrostatic attraction between a shared pair of electrons and the nuclei of the 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.

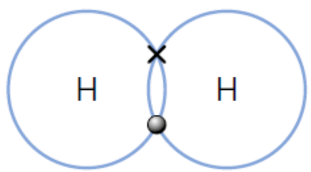
… count the final number of electrons around each atom to make sure that the outer shell is full.

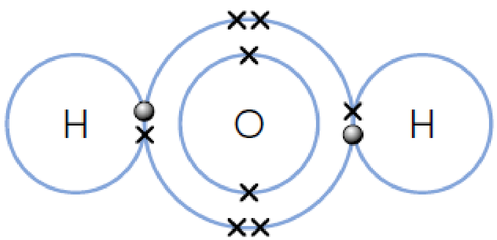
Never …

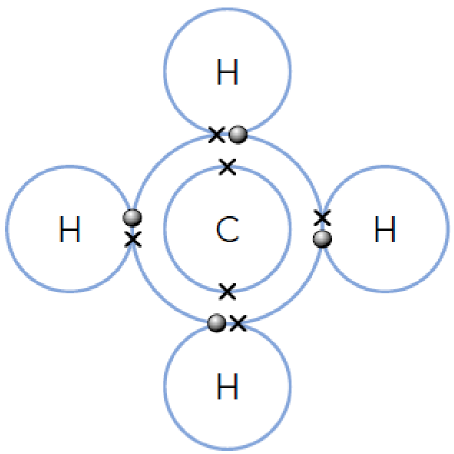
… include a charge on the atoms.

… draw the electron shells separated.

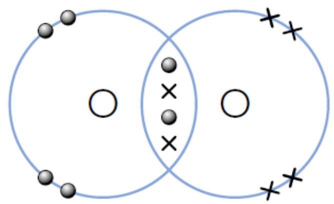
… draw unpaired electrons in the region of overlap.

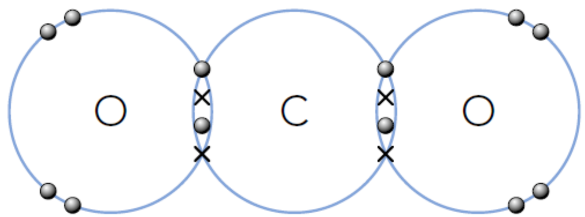






The two diagrams below only show the outer-shell electrons.





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.

(b) Magnesium exists as three stable isotopes. One isotope has a relative isotopic mass of 25.0.

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

(2)

Relative isotopic mass is the mass of an individual atom of a particular isotope relative to 1/12 th the mass of an atom of carbon 12 .

Clearly the candidate recognises that it is the mass of an atom that is being defined and also the relevance of carbon-12.

Example 2 – representations of dot-and-cross diagrams

Perfect answer:

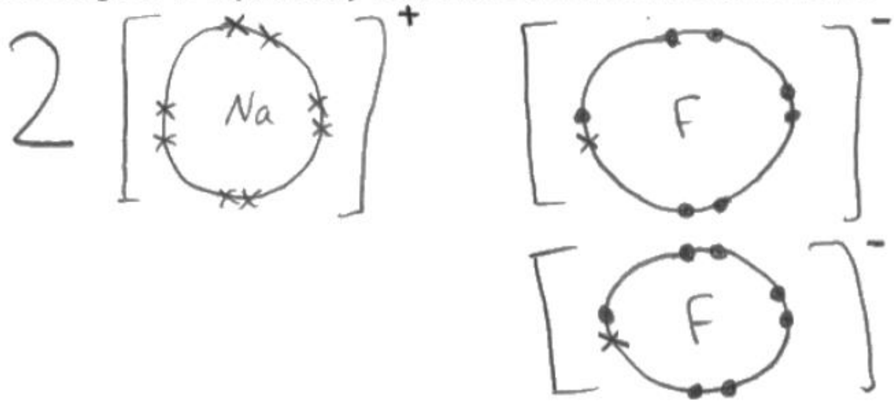
1. Correct charge on BOTH ions.
2. Correct number of outer electrons.
3. No overlap of electron shells – clear separation of ions.

Students need to pay close attention to the instructions given in the question. Some questions on 'dot and cross' diagrams only require the outer electrons to be shown, but some will ask for all the electrons to be shown (usually related to the subsequent part of the question). Here is an example where the candidate lost a mark for not showing 'all' the electrons.

(ii) Draw dot-and-cross diagrams of the ions in sodium fluoride, showing all the electrons.

Use your diagram to explain why the ions are described as isoelectronic.

(3)



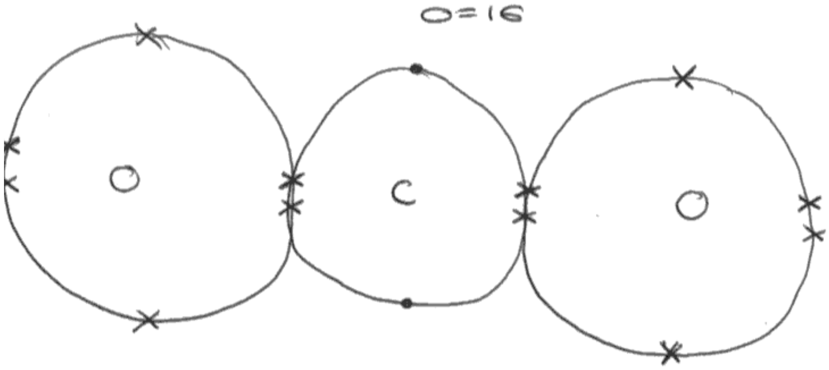
Example 3

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

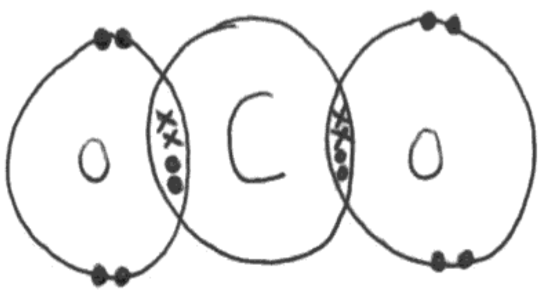
Show outer electrons only.

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

GCSE/GCE Overlap

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)

New concept GCE

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	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$	$\uparrow$	
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?

	1*s*	2*s*	2*p*		
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)

(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.

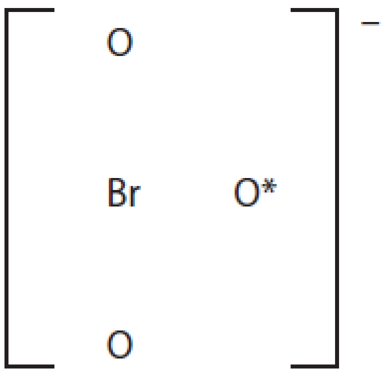
(1 mark)

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

Extra challenge A/A* Booster

14 Potassium bromate(V), KBrO3, is a primary standard, meaning that it can be obtained as a pure substance and used to accurately determine the concentrations of solutions of other chemicals, such as sodium thiosulfate, $Na_2S_2O_3$.

i) Complete the dot and cross diagram for the bromate(V) ion. Show only the outer shell electrons.
In this ion, the bromine expands its outer shell to accommodate 12 electrons.
Use **x** for bromine electrons and • for oxygen electrons. The symbol * on the diagram represents the extra electron which gives the ion its charge.



(2 marks)

ii) Suggest how elements in Period 3 and higher can accommodate more than eight electrons in their outer shell.

(1 mark)

.**(Edexcel GCE, June 2014, WCH02/01,Q19a)**

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 around understanding what is being asked when a practical scenario is given. We have provided a worked example of a question 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. In addition to this, students will need to decide on the appropriate number of significant figures to use in their final answer.

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

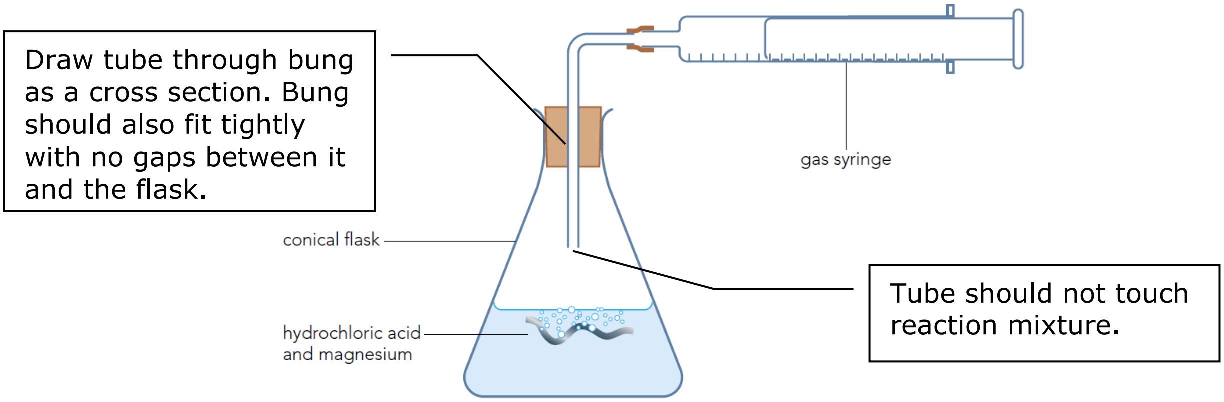
2 Empirical mass =14 + 3 + 32 + 48 = 97

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

b 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 $dm^3\,mol^{-1}$ 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 = ⅔ × 2.08 = 1.39 mol.

4 Convert moles to mass:

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

What is the most appropriate number of significant figures for final answers in calculations?

- Look at the data provided in the question and your final answer should be expressed to the lowest number of significant figures given.
- As a rule, if the question does not state this, typically your answer should be expressed to two or three significant figures. It is rare for answers expressed to one significant figure to score a mark.

Example 1:

This is an example of where the student needs to link all the information given in the stem of the questions and to follow the instructions given. Dilution factor is often missed because students do not follow the sequence of the practical procedures given. It is a good idea to encourage students to 'story board' these or to use flow diagrams to show what is going on.

Titration	1	2	3	4	5
Final burette reading / cm^3	26.00	34.00	36.10	24.15	48.20
Initial burette reading / cm^3	0.00	10.00	11.00	0.05	24.15
Titre / cm^3	26.00	24.00	25.10	24.10	24.05
Concordant results (✓)				✓	✓

(a) Complete the table and determine the concentration, in mol dm^{-3}, of the hydrochloric acid solution, giving the answer to an appropriate number of significant figures.

(5)

mean titre $= \frac{(24.1 + 24.05)}{2} = 24.075$

	HCl	Na_2CO_3
conc	?	
volume	24.075 cm^3	100 cm^3
moles	$0.0123 \times 2 = 0.0246$ moles	$n = m/M_r = \frac{1.3}{106} = 0.0123$ moles

$n = c \times v$

$c = \frac{n}{V} = \frac{0.0246}{(24.075 \div 1000)} = \underline{\underline{1.022 \text{ moldm}^{-3} \text{ of HCl}}}$

This scores two out of five marks:

The initial data is completed correctly, but only two of the three concordant titres are ticked.

The candidate has not divided the number of moles of sodium carbonate by ten to find the number in 10 cm3 so loses a mark here.

The rest of the calculation is correct, but the answer is given to four significant figures which shows greater precision than is possible with the data.

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 mats, 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. Describing metallic structure as 'protons' in a sea of electrons
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.

	What most students can do (well)	Common mistakes
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. Use of terms like 'carbon bonds' instead of covalent bonds

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.

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.

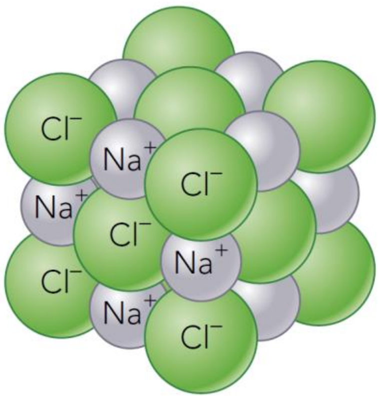
Strength of the metallic bonding increases with the number of valence electrons (outer electrons in the atoms) and with decreasing size of the cation.

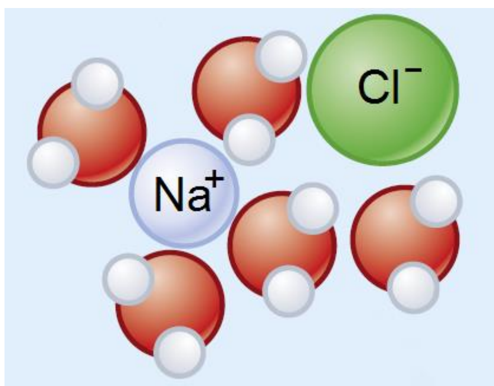
Ionic bonds and ionic compounds

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

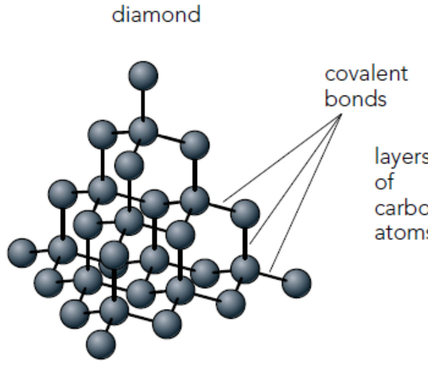
An answer should cover the following points.

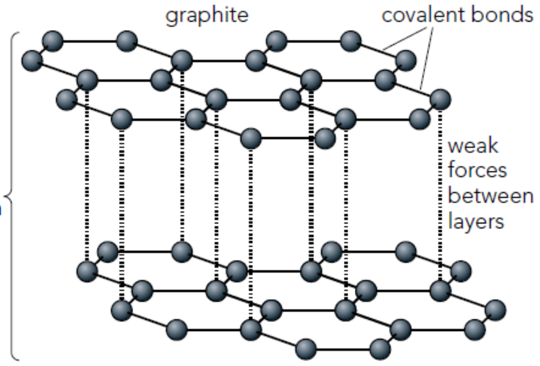
1. The Na^+ and Cl^- ions are held by strong electrostatic forces.
2. To melt solid NaCl, energy is needed to separate overcome the forces of attraction sufficiently for the lattice structure to break down and for the ions to be free to slide past one another.
3. Even though the ions are charged, the 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



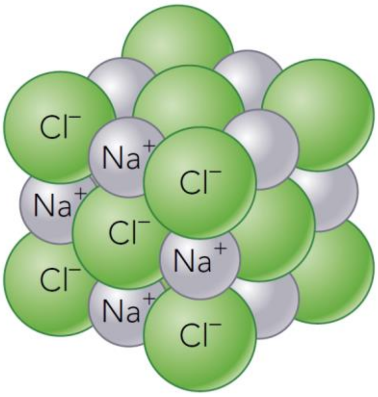


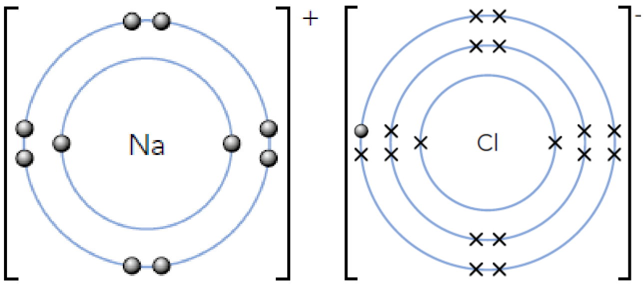
Property	Diamond	Graphite
Melting point	High – atoms held by strong covalent bonds. Many covalent bonds must be broken to melt it. Lots of energy required. Is solid at room temp.	High – atoms held by strong covalent bonds. Many covalent bonds must be broken to melt it. Lots of energy required. Is a solid at room temp.
Electrical conductivity	Poor – no mobile electrons available. All four outer electrons of each carbon are used in bonding.	Good – each carbon only uses three of its outer electron to form covalent bonds. 4th 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.

Examples of common mistakes in exams:

Example 1

1 This question is about covalent bonds.

(a) State what is meant by the term covalent bond.

(2)

A covalent bond is a shared pair of electrons

This level of response will no longer score a mark even at GCSE.

To score two marks, this student needed to mention the nuclei and the attraction between these and the shared pair of electrons, so this scores 0.

Students need to be taught to distinguish between 'how covalent bonds are formed' and 'what is a covalent **bond**'.

Example 2

Mixing up interatomic and intermolecular forces.

NB: at GCSE students will only be required to know general intermolecular forces and the distinction between the different types of intermolecular forces is covered in Topic 2 at GCE.

(b) The strength of ionic bonding in different compounds can be compared by using the amount of energy required to separate the ions. Some values for this energy are given in the table.

Compound	Amount of energy required to separate the ions / $kJ\,mol^{-1}$
LiF	1031
KF	817
CaF_2	2957

Using the data provided, explain how changes in the cation affect the bond strength in an ionic compound.

(2)

The larger the cation the weaker the bond as the attraction between the nuclear charge of the cation and the anion is decreased. Additionally, the charge of the cation affects strength as the higher the charge the greater the charge density, so there's stronger attraction between oppositley charged ions (eg Ca^{2+} not K^+).

This student clearly knows and understands the factors which will result in different amounts of energy being required to separate the ions.

Unfortunately, the question requires the discussion to be in terms of the data provided for full marks to be awarded.

This has not happened here so this scores 1 mark.

Example 3

(a) The boiling temperatures of three halogens are shown in the table.

Halogen	Boiling temperature / °C
chlorine	−35
bromine	59
iodine	184

Explain why the boiling temperatures increase from chlorine to iodine.

Explain why the boiling temperatures increase from chlorine to iodine.

(2)

The boiling temperature increases because iodine is such a large atom that it has powerful london forces holding the atoms together holding it requires more energy in the form of heat to break down

This candidate thinks that London forces are holding the atoms together and that more energy is needed to break them down. This implies that the bonds between the atoms are breaking, which is incorrect so this response scored 0.

A better response:

(2)

- Atomic radius of the halogens increases
- Therefore intermolecular forces (London forces) become stronger and more energy is required to break them (and boil the diatomic halogens)

This answer scored the second mark in the mark scheme. Increasing atomic radius is not sufficient for explaining the increasing strength of the London forces.

The best response:

(2)

Because there are more ~~[illegible]~~ electrons as you go down group 7 so there are more and stronger London forces so more energy is required to break the London forces as you go down group 7

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)

Edexcel GCE Jan 2011, 6CH01, Q17

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)

Edexcel GCE Jan 2012, 6CH01

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

Useful concepts and contents better supported by the GCSE 2016 specification compared to previous specifications:

1 selecting the correct number of significant figures to express their final answer and giving correct units

2 structure 6-mark questions to give a balance of 'facts' and linked statements

3 data analysis and evaluation – refer to data given when answering questions

4 definition of bonds in terms of electrostatic forces

5 answering unscaffolded calculation questions

6 titration is now covered in both Combined and Separate Chemistry

7 reacting masses – both Combined and Separate Chemistry students should be able to evaluate data to determine which is the limiting reagent

8 writing ionic equations

9 relationship between concentration of H+ ions and pH and effect of dilution

10 distinction between strengths of acids and concentration.

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	Topic1: Key Concept SC3/CC3: Atomic Structure 1.2 Describe the structure of an atom as a nucleus containing protons and neutrons, surrounded by electrons in shells
2. know the relative mass and relative charge of protons, neutrons and electrons	1.3 Recall the relative charge and relative mass of: a a proton b a neutron c an electron 1.4 Explain why atoms contain equal numbers of protons and electrons 1.5 Describe the nucleus of an atom as very small compared to the overall size of the atom 1.6 Recall that most of the mass of an atom is concentrated in the nucleus 1.7 Recall the meaning of the term mass number of an atom 1.8 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
3. know what is meant by the terms 'atomic (proton) number' and 'mass number'	1.7 Recall the meaning of the term mass number of an atom

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	1.10 Calculate the numbers of protons, neutrons and electrons in atoms given the atomic number and mass number 1.23 Calculate the numbers of protons, neutrons and electrons in simple ions given the atomic number and mass number 1.11 Explain how the existence of isotopes results in relative atomic masses of some elements not being whole numbers
5. understand the term 'isotopes'	1.9 Describe isotopes as different atoms of the same element containing the same number of protons but different numbers of neutrons in their nuclei
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.*	
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*	1.12 **Calculate the relative atomic mass of an element from the relative masses and abundances of its isotopes**

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 21. 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 22. know that elements can be classified as *s*-, *p*- and *d*-block elements	SC4/CC4: The Periodic Table 1.19 Predict the electronic configurations of the first 20 elements in the periodic table as diagrams and in the form, for example 2.8.1 1.20 Explain how the electronic configuration of an element is related to its position in the periodic table
Topic 2 – Bonding and structure	**GCSE**
1. know that ionic bonding is the strong electrostatic attraction between oppositely charged ions	
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	SC5/CC5:Ionic Bonds 1.21 Explain how ionic bonds are formed by the transfer of electrons between atoms to produce cations and anions, including the use of dot and cross diagrams 1.22 Recall that an ion is an atom or group of atoms with a positive or negative charge 1.27 Explain the structure of an ionic compound as a lattice structure a consisting of a regular arrangement of ions b held together by strong electrostatic forces (ionic bonds) between oppositely charged ions

7. know that a covalent bond is the strong electrostatic attraction between two nuclei and the shared pair of electrons between them	SC6/CC6: Covalent Bonds 1.28 Explain how a covalent bond is formed when a pair of electrons is shared between two atoms 1.29 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	1.31 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
22. know that metallic bonding is the strong electrostatic attraction between metal ions and the sea of delocalised electrons	SC7/CC7: Types of substances and metallic bonds 1.40 Explain the properties of metals, including malleability and the ability to conduct electricity 1.42 Describe most metals as shiny solids which have high melting points, high density and are good conductors of electricity whereas most non-metals have low boiling points and are poor conductors of electricity

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	1.32 Explain why elements and compounds can be classified as: a ionic b simple molecular (covalent) c giant covalent d metallic and how the structure and bonding of these types of substances results in different physical properties, including relative melting point and boiling point, relative solubility in water and ability to conduct electricity (as solids and in solution) 1.35 Recall that graphite and diamond are different forms of carbon and that they are examples of giant covalent substances 1.36 Describe the structures of graphite and diamond 1.38 Explain the properties of fullerenes including C60 and graphene in terms of their structures and bonding
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	
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	SC7/CC7: Types of Substances 1.33 Explain the properties of ionic compounds limited to: a high melting points and boiling points, in terms of forces between ions b whether or not they conduct electricity as solids, when molten and in aqueous solution 1.34 Explain the properties of typical covalent, simple molecular compounds limited to: a low melting points and boiling points, in terms of forces

	between molecules (intermolecular forces) 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	SC9/CC9: Calculations with masses **1.50 Recall that one mole of particles of a substance is defined as:** **a the Avogadro constant number of particles (6.02×10^{23} atoms, molecules, formulae or ions) of that substance** **b a mass of 'relative particle mass' g**
2. be able to use the Avogadro constant, *L*, (6.02×10^{23} mol^{-1}), in calculations	
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'	1.45 Deduce: a the empirical formula of a compound from the formula of its molecule b the molecular formula of a compound from its empirical formula and its relative molecular mass
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 -and ions 0.2 Write word equations 0.3 Write balanced chemical equations including the use of state symbols (s), (l), (g) and (aq) 0.4 Write balanced ionic equations
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.*	SC9/CC9: Calculations with masses 1.47 Explain the law of conservation of mass applied to: a a closed system including a precipitation reaction in a closed flask b a non-enclosed system including a reaction in an open flask that takes in or gives out a gas 1.48 Calculate masses of reactants and products from balanced equations, given the mass of one substance 1.51 **Calculate the number of:** **a moles of particles of a substance in a given mass of that substance and vice versa** **b particles of a substance in a given number of moles of that substance and vice versa** **c particles of a substance in a given mass of that substance and vice versa** 1.52 **Explain why, in a reaction, the mass of product formed is controlled by the mass of the reactant which is not in excess** 1.53 **Deduce the stoichiometry of a reaction from the masses of the reactants and products**
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***	
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	Topic 5 – SC14 5.8C **Calculate the concentration of solutions in mol dm^{-3} and convert concentration in g dm^{-3} into mol dm^{-3} and vice versa** 5.9C *Core Practical: Carry out an accurate acid-alkali titration, using burette, pipette and a suitable indicator* 5.10C **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 = $\frac{\text{(molar mass of the desired product)}}{\text{(sum of the molar masses of all products)}} \times 100\%$	5.11C Calculate the percentage yield of a reaction from the actual yield and the theoretical yield 5.12C Describe that the actual yield of a reaction is usually less than the theoretical yield and that the causes of this include: a incomplete reactions b practical losses during the experiment c competing, unwanted reactions (side reactions) 5.13C Recall the atom economy of a reaction forming a desired product 5.14C Calculate the atom economy of a reaction forming a desired product

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.	Topic 3 – SC8/CC8: Acids 3.19 Recall the general rules which describe the solubility of common types of substances in water: a all common sodium, potassium and ammonium salts are soluble b all nitrates are soluble c common chlorides are soluble except those of silver and lead d common sulfates are soluble except those of lead, barium and calcium e common carbonates and hydroxides are insoluble except those of sodium, potassium and ammonium 3.11 Explain the general reactions of aqueous solutions of acids with: a metals b metal oxides c metal hydroxides d metal carbonates to produce salts Topic4 – SC11/CC11: Obtaining and using metals 4.1 Deduce the relative reactivity of some metals, by their reactions with water, acids and salt solutions 4.2 **Explain displacement reactions as redox reactions, in terms of gain or loss of electrons** 4.3 Explain the reactivity series of metals (potassium, sodium, calcium, magnesium, aluminium, (carbon), zinc, iron, (hydrogen), copper, silver, gold) in terms of the reactivity of the metals with water and dilute acids and that these reactions show the relative tendency of metal atoms to form cations

Appendix 2: Answers to Baseline assessment

1 Copper(II) sulfate – $CuSO_4$
Sodium hydroxide – NaOH
Strontium nitrate – $Sr(NO_3)_2$
Sodium carbonate – Na_2CO_3
Lithium hydrogencarbonate – $LiHCO_3$
Potassium nitrate – KNO_3
Calcium hydroxide – $Ca(OH)_2$
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
C_2H_5OH – 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 × 16.0) = 64.1 |
| C_2H_6 | (2 × 12.0) + (6 × 1.0) = 30.0 |
| C_2H_5OH | (2 × 12.0) + (5 × 1.0) + 16 + 1 = 46.0 |
| NH_4Cl | 14.0 + (4 × 1.0) + 35.5 = 53.5 |
| KBr | 39.1 + 79.9 = 119 |
| $Ca(OH)_2$ | 40.1 +(2 × (1.0 + 16.0))=74.1 |
| $NaNO_3$ | 23.0 + 14.0 + (3 × 16.0) = 85.0 |
| $FeCl_3$ | 55.8 + (35.5 × 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

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

Magnesium and sulfuric acid

magnesium and sulfuric acid → magnesium sulfate + hydrogen

$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

Complete combustion of butane

butane + oxygen → carbon dioxide + water

$C_4H_{10}(g) + 6\frac{1}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$

OR

$2C_4H10(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$

Thermal decomposition of calcium carbonate

calcium carbonate → calcium oxide + carbon dioxide

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

Sodium and water

sodium + water → sodium hydroxide + hydrogen gas

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH_{(aq)} + H_2(g)$

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	H_2S	**covalent**	**(simple) molecular**
Graphite	**C**	**covalent**	**giant**
Silicon dioxide	SiO_2	**covalent**	**giant**
Methane	CH_4	**covalent**	**(simple) molecular**
Calcium	**Ca**	**metallic**	**giant**
Magnesium chloride	$MgCl_2$	**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.

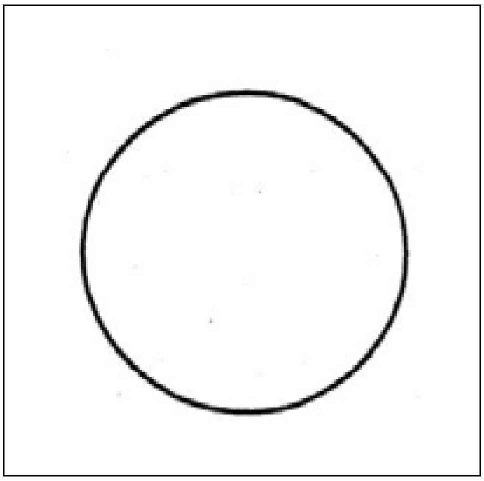
For positive ions this decreases by the number equivalent to the charge on the ion.

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

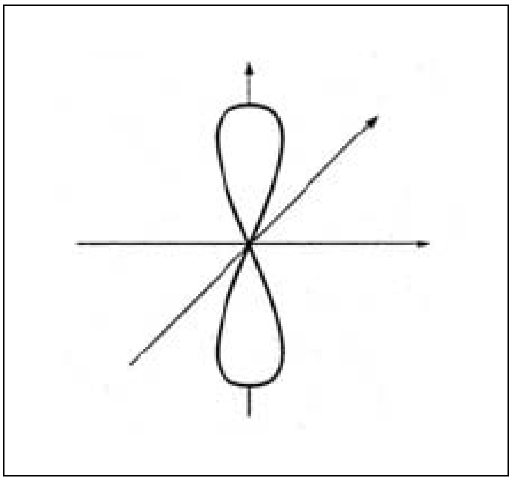
Worksheet 2 answers

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**	**$2s^2sp^6$**
n = 3	**18**	**s, p, d**	**9**	**$3s^23p^63d^{10}$**
n = 4	**32**	**s, p ,d, f**	**16**	**$4s^24p^64d^{10}4f^{14}$**

s



p



		Electron configuration		Electrons in boxes														
	Z	**2.8.8**	**s,p,d**	**1s**	**2s**	**2p**			**3s**	**3p**			**3d**					**4s**
Na	11	2.8.1	$1s^22s^2sp^63s^1$	↑↓	↑↓	↑↓	↑↓	↑↓	↑									
Be	4	2.2	$1s^22s^2$	↑↓	↑↓													
Be^{2+}	4	2	$1s^2$	↑↓														
P	15	2.8.5	$1s^22s^22p^63s^23p^3$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑						
Cr	24	2.8.13.1	$1s^22s^22p^63s^23p^63d^54s^1$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑	↑	↑
Cu	29	2.8.18.1	$1s^22s^22p^63s^23p^63d^{10}4s^1$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑
Fe	26	2.8.14.2	$1s^22s^22p^63s^23p^63d^64s^2$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑	↑↓
Al	13	2.8.3	$1s^22s^22p^63s^23p^1$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑								
Al^{3+}	13	2.8	$1s^22s^22p^6$	↑↓	↑↓	↑↓	↑↓	↑↓										
Sc	21	2.8.9.2	$1s^22s^22p^63s^23p^63d^14s^2$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑					↑↓
Cl	17	2.8.7	$1s^22s^22p^63s^23p^5$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑						
Cl^-	17	2.8.8	$1s^22s^22p^63s^23p^6$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓						

Section B

Worksheet 1 answers

Copper(II) sulfate	**$CuSO_4$**
Nitric acid	**HNO_3**
Copper(II) nitrate	**$Cu(NO_3)_2$**
Sulfuric acid	**H_2SO_4**
Sodium carbonate	**Na_2CO_3**
Aluminium sulfate	**$Al_2(SO_4)_3$**
Ammonium nitrate	**NH_4NO_3**
Nitrogen dioxide	**NO_2**
Sulfur dioxide	**SO_2**
Ammonia	**NH_3**
Ammonium sulfate	**$(NH_4)_2SO_4$**
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	**$AgNO_3$**	**Ag^+**	**NO_3^-**
Copper(II) sulfate	**$CuSO_4$**	**Cu^{2+}**	**SO_4^{2-}**
Sodium hydrogencarbonate	$NaHCO_3$	**Na^+**	**HCO_3^-**
Magnesium carbonate	**$MgCO_3$**	**Mg^{2+}**	**CO_3^{2-}**
Lithium carbonate	**Li_2CO_3**	**$2Li^+$**	**CO_3^{2-}**
Calcium hydrogensulfate	$Ca(HSO_4)_2$	**Ca^{2+}**	**HSO_4^-**
Aluminium nitrate	**$Al(NO_3)$**	**Al^{3+}**	**$3NO_3^-$**
Calcium phosphate	**$Ca_3(PO_4)_2$**	**$3Ca^{2+}$**	**$2PO_4^{3-}$**
Potassium hydride	**KH**	**K^+**	**H^-**
Sodium ethanoate	**CH_3COONa**	**Na^+**	**CH_3COO^-**
Potassium manganate(VII)	$KMnO_4$	**K^+**	**MnO_4^-**
Potassium dichromate(VI)	**$K_2Cr_2O_7$**	**$2K^+$**	**$Cr_2O_7^{2-}$**
Zinc chloride	**$ZnCl_2$**	**Zn^{2+}**	**$2Cl^-$**

Substance	Formula	Cation (exact number)	Anion (exact number)
Strontium nitrate	**$Sr(NO_3)_2$**	**Sr^{2+}**	**$2NO_3^-$**
Sodium chromate(VI)	**$NaCrO_4$**	**Na^+**	**CrO_4^{2-}**
Calcium fluoride	**CaF_2**	**Ca^{2+}**	**$2F^-$**
Potassium sulfide	**K_2S**	**$2K^+$**	**S^{2-}**
Magnesium nitride	**Mg_3N_2**	**$3Mg^{2+}$**	**$2N^{3-}$**
Lithium hydrogensulfate	**$Li(HSO_4)$**	**Li^+**	**HSO_4^-**
Ammonium sulfate	$(NH_4)_2SO_4$	**$2NH_4^+$**	**SO_4^{2-}**

Worksheet 3 answers

1 $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
$Mg(aq) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

2 $CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l)$
$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$

3 $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

4 $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$
$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

5 $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$
$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

6 $2AgNO_3(aq) + MgCl_2(aq) \rightarrow 2AgCl(s) + Mg(NO_3)_2(aq)$
$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

7 $MgO(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2O(l)$
$MgO(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2O(l)$

8 $CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$
$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$

9 $Pb(NO_3)(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$
$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$

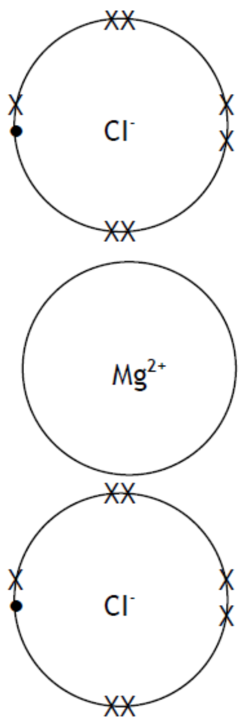
10 $Fe_2(SO_4)_3 + 6NaOH(aq) \rightarrow 3Fe(OH)_3(s) + 3Na_2SO_4(aq)$
$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

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 Mg^{2+} 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 CO_2 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 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 $\frac{1}{12}$ to 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 × 10) + (80.3 × 11) (1 mark) ÷ 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.

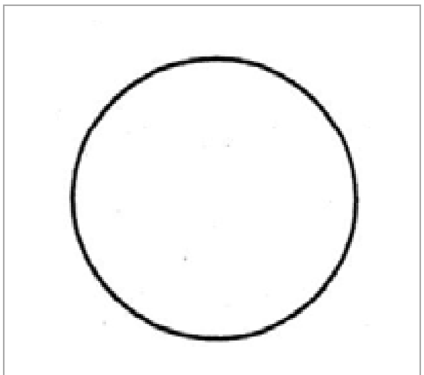
		4s	4p		
As	[Ar] $3d^{10}$	↑↓	↑	↑	↑
Se	[Ar] $3d^{10}$	↑↓	↑↓	↑	↑

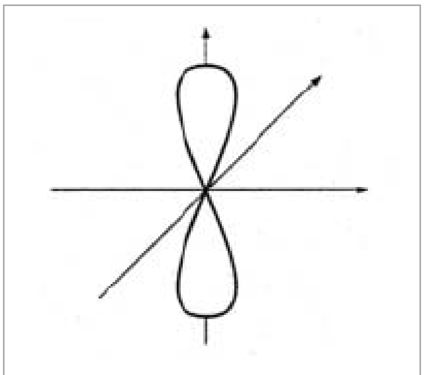
b i A region (around the nucleus) where there is a high probability of finding an electron.

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.


s-orbital


p-orbital

c 11/eleven (1 mark)

d 18/eighteen (1 mark)

12 a (194 × 32.8) + (195 × 30.6) + (53.0 × 25.4) +(198 × 11.2) ÷ 100 (1 mark)
= 195.262
= 195.3 (1 dp) (1 mark)

b d(-block) (1 mark)

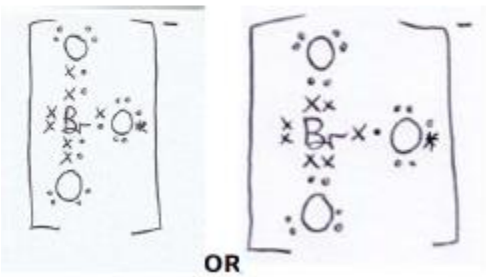
13 a 1 mark for each correct row.

Isotope	$^{131}_{53}I$	$^{127}_{53}I$
Number of protons	**53**	**53**
Number of neutrons	**78**	**74**

(2 marks)

b Xenon/Xe (1 mark)

14
(i)



7x and 5 • around the bromine (1)
Total of 8 electrons round each oxygen and one octet MUST INCLUDE the electron represented by * (1)

ALLOW

x for oxygen and • for bromine if clear

Electrons in bonds to be shown in rows eg xx •• or x•x• between the relevant atoms; non-bonded electrons not in pairs.

All dots or all crosses then max 1

Two dative covalent bonds by the bromine to the oxygens then max 1 (loses first mark)

IGNORE

circles round outer shells of atoms

(ii) There are vacant (3)**d** orbitals / they are using (3)**d** orbitals (1)

ALLOW

Sub-shells for orbitals

Use of D for d

DO NOT ALLOW

2d, p/ f orbitals

Shell for subshell

Section B

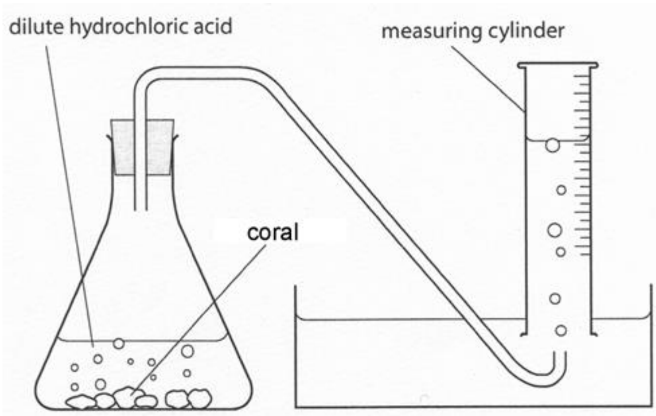
1 **a** $H_2O + CO_2 \rightarrow H_2CO_3$ (1 mark)

b $2H^+ + CO_3^{2-}$ (1 mark) $\rightarrow H_2O + CO_2$ (1 mark)

OR

$2H_3O^+ + CO_3^{2-}$ (1 mark) $\rightarrow 3H_2O + CO_2$ (1 mark)

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 ÷ 24000 =) 0.00933 = 9.33×10^{-3} (mol) (1 mark).

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

Reject '0.009' as answer.

d $CaCO_3(s) + 2HCl(aq.) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$ (1 mark)
All four state symbols must be correct for this mark

e Mass of 1 mol $CaCO_3$ = 40.1 + 12.0 + (3 × 16.0) = 100.1 g (1 mark).

f Mass of $CaCO_3$ = 100.1 × 0.00933 = 0.934 (g) (1 mark)
Percentage of $CaCO_3$ in the coral = 100 x 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).

3 **a** $MgCO_3(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l) + CO_2(g)$
OR
$MgCO_3(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + CO_2(g) + H_2O(l)$
All formulae and balancing (1 mark).
Mark state symbols independently (1 mark)

b $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$
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 $MgCl_2$) / stronger ionic bonding in MgO (than in $MgCl_2$) (1 mark)

therefore, more energy is required to overcome the forces of attraction in MgO (than in $MgCl_2$) / more energy is required to break the (ionic) bonds in MgO (than in $MgCl_2$). (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				
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

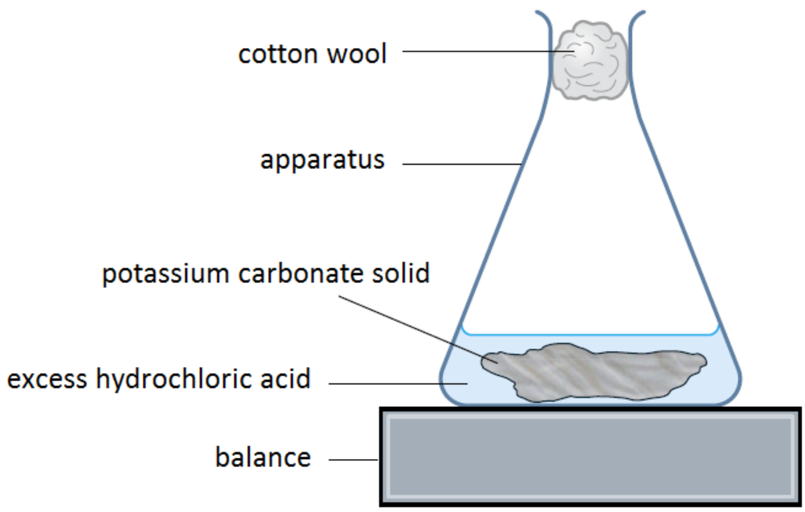
(18 marks)

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 four 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)

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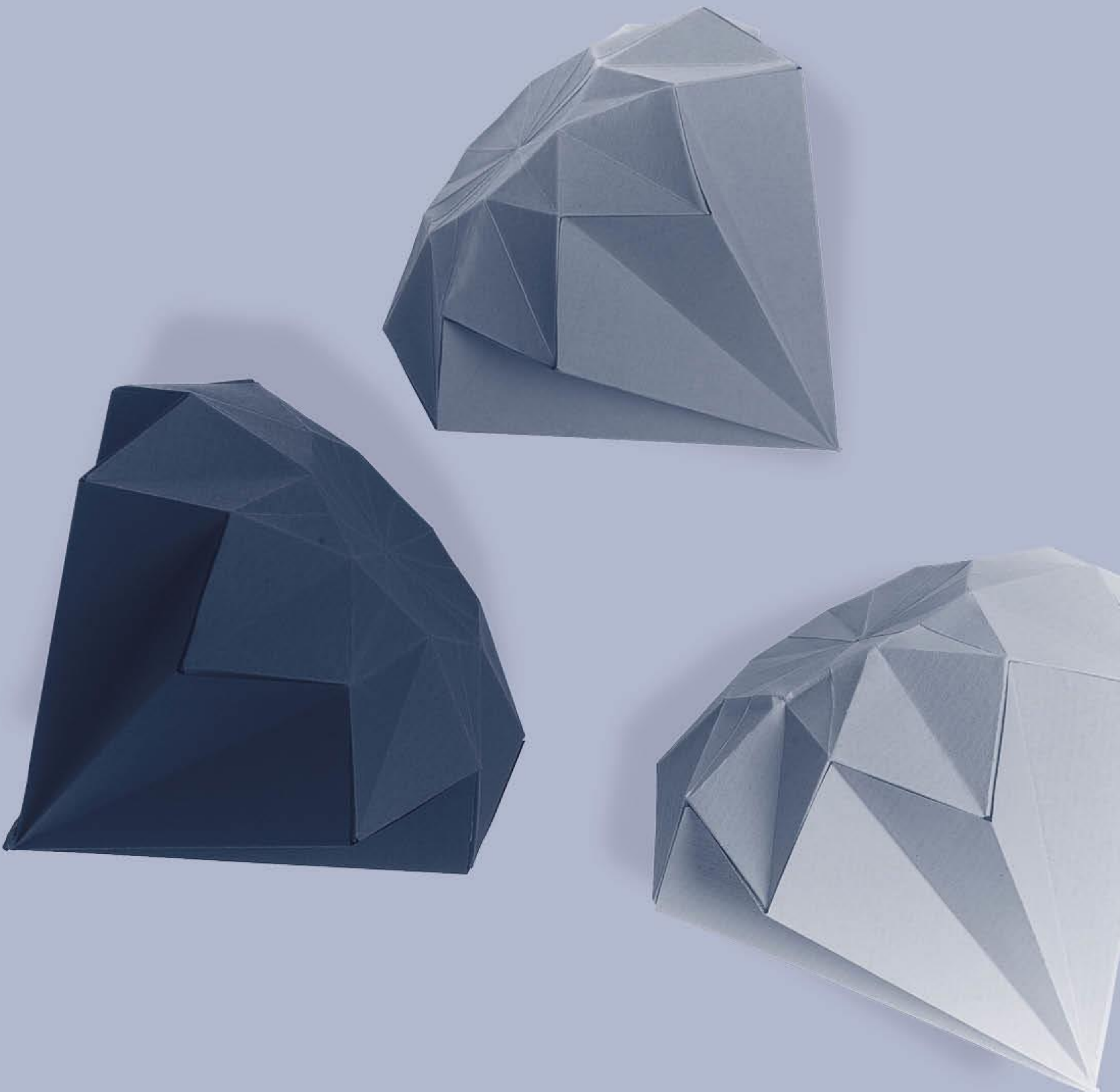

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