The Henry Box School

Chemistry – Preparation for A Level

Acceptance criteria: 6 in GCSE Chemistry or 6-6 in Combined Science

6 in GCSE Maths

Course: Edexcel Chemistry (2015)

A Level Chemistry covers a very broad range of topics and involves the use of many scientific skills. To cover these topics in more detail, it is essential that all students have a confident grasp of related work covered in GCSE Chemistry or Combined Science.

Therefore, we require that all students **complete this booklet** of review tasks that will help to refresh your knowledge and bridge the gap between GCSE and A Level Chemistry.

You may find that some of the tasks go beyond the scope of the GCSE Chemistry course you studied. This is A Level content that has been included to stretch and challenge you. Use the guide sheets provided to help you.

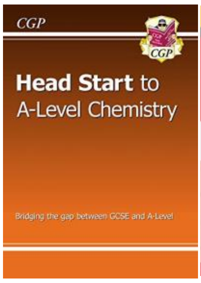
All students will sit a **baseline assessment** in the first A Level Chemistry lesson in September which will be based on the work covered in this booklet from your GCSE.

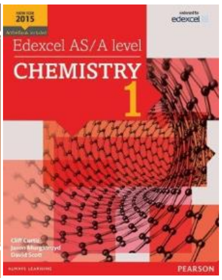
You should bring the completed booklet to **your first chemistry lesson**.

You may also wish to work through Headstart to A Level Chemistry, which is availably to download for free on Kindle currently or about £5 otherwise.

The course textbook for September is Edexcel AS/A level Chemistry Student Book 1
ISBN 9781447991168

Have a great summer – it's been a tough one – and we look forward to seeing you in September!





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Section A: Forming ions and subatomic particles

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*.

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

Formation of ions

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

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

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

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

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

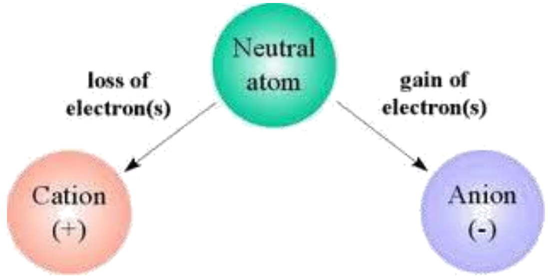
Atoms of Group 5 non-metals gain three electrons and form ions with a 3– charge, e.g. N^{3-}

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

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

ANions = Negative

Ca+ions = +ive



Why are ions negative or positive?

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

KS4 – Electron configuration

Filling electron shells

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

Representing electron configurations

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

Using the Periodic Table

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

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

KS4 – Dot-and-cross diagrams for ionic bonding

Hints and tips

Always …

… count the electrons!

… remember that ions should have full outer shells.

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

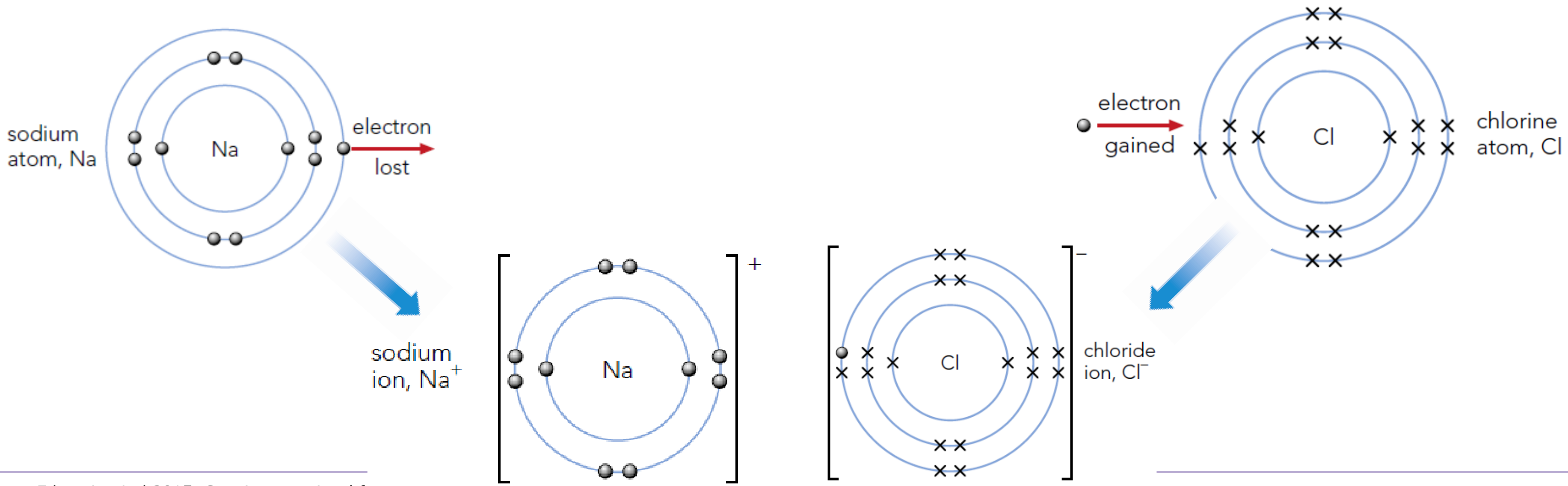
Never …

… show the electron shells overlapping.

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

… remove electrons from the inner shell.

… give metals a negative charge.



KS4 – Covalent compounds (simple covalent bonding)

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

Covalent bonding results in the formation of molecules.

Hints and tips

Always …

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

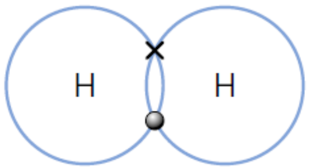
… count the final number of electrons around each atom to make sure that the outer shell is full.

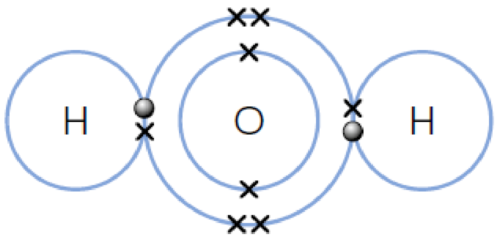
Never …

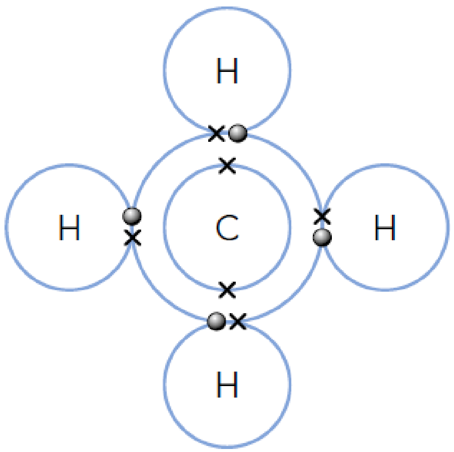
… include a charge on the atoms.

… draw the electron shells separated.

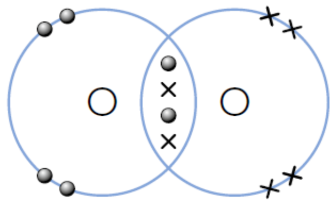
… draw unpaired electrons in the region of overlap.

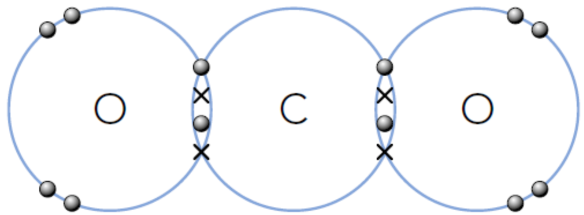






The two diagrams below only show the outer-shell electrons.





Worksheet 1: Atomic structure and the Periodic Table

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

Structure of an atom

The nucleus contains …

The electrons are found in the …

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

In a neutral atom the number of electrons is …

To work out the number of neutrons we …

Vocabulary

State what is meant by the following terms.

1 Relative atomic mass

2 Relative molecular mass

3 Isotope

4 Relative isotopic mass

Structure of an ion

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

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

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

Section B: Writing Formulae

Summary sheet

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

Worked examples: Calculations

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

Empirical formulae

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

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

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

Interpreting the question

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

Answering the question

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

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

Answering the question

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

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

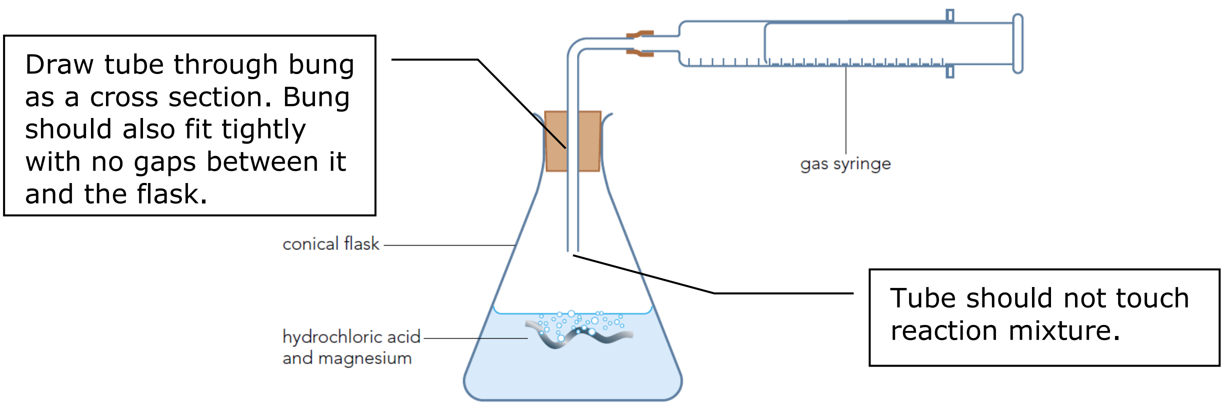
2 Empirical mass =14 + 3 + 32 + 48 = 97

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

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

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

Answering the question



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

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 hydroxid

Section C: Structure and properties

Summary sheet 1: Structure and bonding

Words used to describe structure and bonding:

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

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

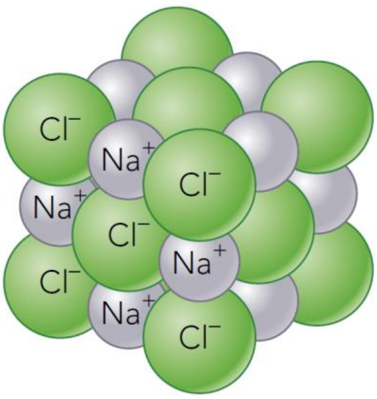
Strength of the metallic bonding increases with the number of valence electrons (outer electrons in the atoms) and with decreasing size of the cation.

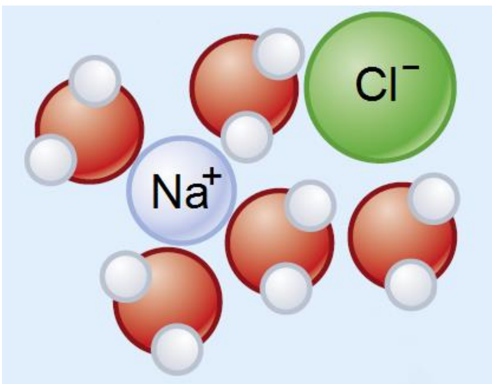
Ionic bonds and ionic compounds

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

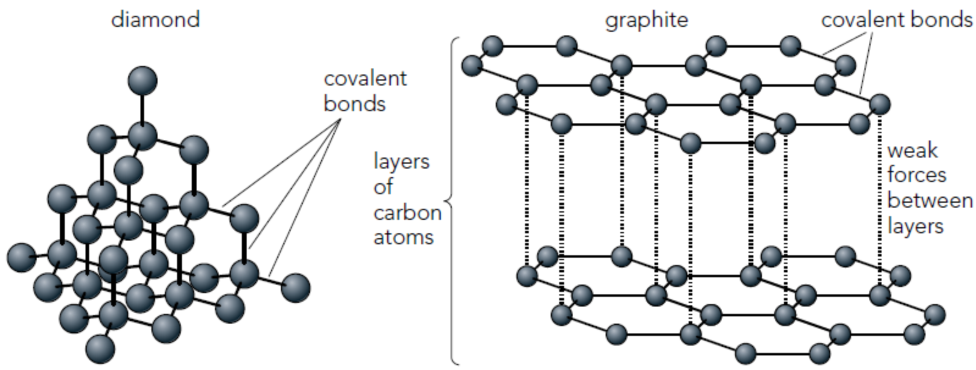
An answer should cover the following points.

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





Summary sheet 2: Diamond and graphite structures

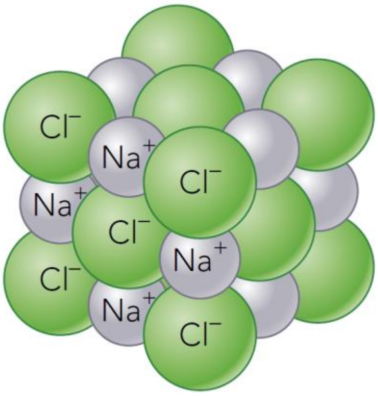


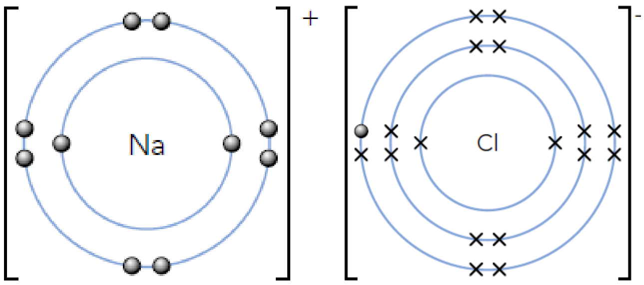
Property	Diamond	Graphite
Melting point	High – atoms held by strong covalent bonds. Many covalent bonds must be broken to melt it. Is solid at room temp.	High – atoms held by strong covalent bonds. Many covalent bonds must be broken to melt it. Is a solid at room temp.
Electrical conductivity	Poor – no mobile electrons available. All 4 outer electrons of each carbon are used in bonding.	Good – each carbon only uses 3 of its outer electron to form covalent bonds. 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.

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)