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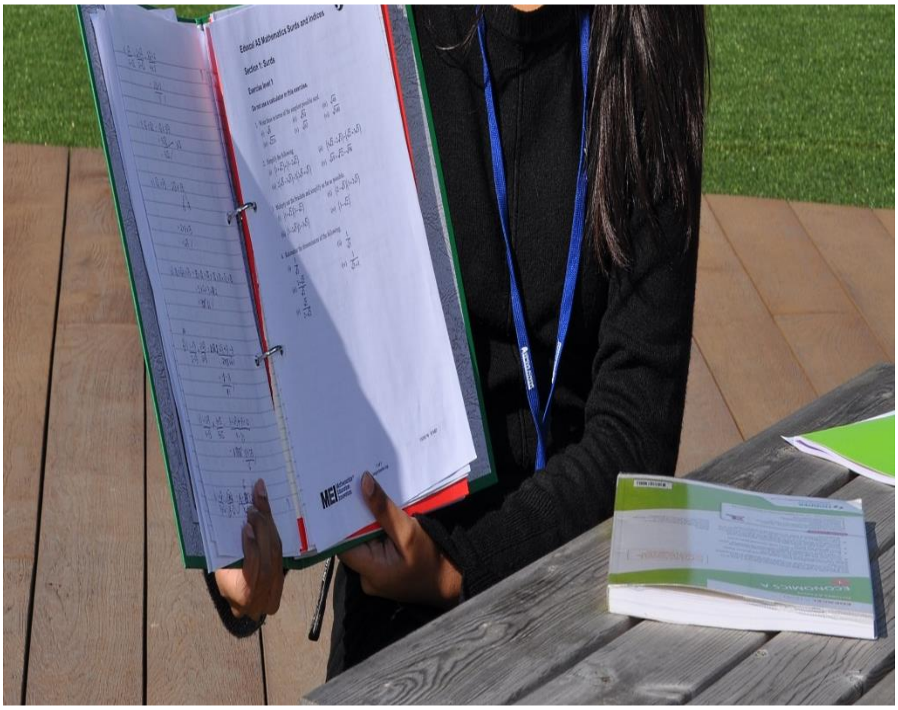


LAE Tottenham Preparatory Tasks for Offer Holders

Preparatory Tasks: Chemistry

Welcome to A-Level Chemistry!

We hope you are looking forward to undertaking your A Level Chemistry course.

Moving from GCSE Science to A Level can be a daunting leap. You'll be expected to remember a lot more facts, equations, and definitions, and you will need to learn new maths skills and develop confidence in applying what you already know to unfamiliar situations.

You will have two A Level Chemistry teachers. One will look at atomic and electronic structure and then structure and bonding and the other will start with quantitative chemistry beginning with moles.

In advance of your first lessons in September, we would like you to do three things:

Task 1: Math Skills Practice (5 hours)

Use the checklist provided and work through the questions in appendix 1. Some of the chemistry concepts might be new to you, read through the examples and attempt as many questions as you can for each section. Bring evidence of this work on your first day and ready to go into your folder. Answers at the end.

Amount of Substance	
	I can convert units and standard form and decimals (Section 1)
	I can balance equations (Section 2)
	I can rearrange equations (Section 3)
	I can carry out calculations using concentration, volume and amount of substance in a solution
	I can calculate moles and masses (Section 4)
	I can carry out calculations using mass of substance, Mr and amount in moles
	I can calculate percentage yield and atom economy from given data (Section 5)
	I can calculate percentage error
	I can define relative atomic mass (Ar) and relative molecular mass (Mr) (Section 6)
	I can carry out calculations using the Avogadro constant
	I can calculate empirical formula from data giving composition by mass or percentage mass (Section 7)
	I can calculate molecular formula from the empirical formula and relative molecular mass
	I can calculate the concentration of a solution from mean titre results

Task 2: GCSE Chemistry topics - (3 hours)

Using the checklist provided revisit a few topics from GCSE Chemistry: you will need to produce fresh, neat, well organised, illustrated, sensibly coloured notes on your prior understanding of the topics in the checklist on A4 paper, ready to go into your folder, to be submitted to your teacher in your first lesson. See appendix 2 for brief notes and questions to consider. We will be especially impressed if you have considered the A Level content coming in those first lessons. We will not be impressed if you turn up with your old unimproved GCSE materials.

Structure and Bonding	
	I can describe ionic bonding
	I can predict the charge on a simple ion using the position of the element in the periodic table
	I can construct formulas for ionic compounds (Mg^{2+} and Cl^- becomes $MgCl_2$)
	I can describe the difference between a single covalent bond and a co-ordinate (dative covalent) bond
	I can represent a covalent bond using a line and a co-ordinate bond using an arrow
	I can describe metallic bonding
	I can describe and explain the properties of: diamond, graphite, iodine, magnesium and sodium chloride as examples of one of these 4 crystal structures: ionic, metallic, Giant Covalent, simple molecular
	I can relate the melting point and conductivity of materials to the type of structure and bonding present

Task 3: A Level Definitions (3 hours)

The aim of this section is to pre-learn some useful knowledge from the first chapters of your A Level course. This can be found in appendix 3.

Appendix 1 Maths Skills

1 Core mathematical skills

A practical chemist must be proficient in standard form, significant figures, decimal places, SI units, and unit conversion.

1.1 Standard form

In science, very large and very small numbers are usually written in standard form. Standard form is writing a number in the format $A \times 10^x$ where A is a number from 1 to 10 and x is the number of places you move the decimal place.

For example, to express a large number such as 50 000 mol dm^{-3} in standard form, A = 5 and x = 4 as there are four numbers after the initial 5.

Therefore, it would be written as 5×10^4 mol dm^{-3}.

To give a small number such as 0.000 02 Nm^2 in standard form, A = 2 and there are five numbers before it so x = −5.

So it is written as 2×10^{-5} Nm^2.

Practice questions

1 Change the following values to standard form.

a. boiling point of sodium chloride: 1413 °C

b. largest nanoparticles: 0.0001×10^{-3} m

c. number of atoms in 1 mol of water: 1806×10^{21}

2 Change the following values to ordinary numbers.

a. 5.5×10^{-6} **b.** 2.9×10^2 **c.** 1.115×10^4 **d.** 1.412×10^{-3} **e.** 7.2×10^1

1.2 Significant figures and decimal places

In chemistry, you are often asked to express numbers to either three or four significant figures. The word significant means to ‘have meaning’. A number that is expressed in significant figures will only have digits that are important to the number’s precision.

It is important to record your data and your answers to calculations to a reasonable number of significant figures. Too many and your answer is claiming an accuracy that it does not have, too few and you are not showing the precision and care required in scientific analysis.

For example, 6.9301 becomes 6.93 if written to three significant figures.

Likewise, 0.000 434 56 is 0.000 435 to three significant figures.

Notice that the zeros before the figure are *not* significant – they just show you how large the number is by the position of the decimal point. Here, a 5 follows the last significant digit, so just as with decimals, it must be rounded up.

Any zeros between the other significant figures are significant. For example, 0.003 018 is 0.003 02 to three significant figures.

Sometimes numbers are expressed to a number of decimal places. The decimal point is a place holder and the number of digits afterwards is the number of decimal places.

For example, the mathematical number pi is 3 to zero decimal places, 3.1 to one decimal place, 3.14 to two decimal places, and 3.142 to three decimal places.

Practice questions

3 Give the following values in the stated number of significant figures (s.f.).

a 36.937 (3 s.f.) **b** 258 (2 s.f.) **c** 0.043 19 (2 s.f.) **d** 7 999 032 (1 s.f.)

4 Use the equation:

number of molecules = number of moles × 6.02 × 10^{23} molecules per mole to calculate the number of molecules in 0.5 moles of oxygen. Write your answer in standard form to 3 s.f.

5 Give the following values in the stated number of decimal places (d.p.).

a 4.763 (1 d.p.) **b** 0.543 (2 d.p.) **c** 1.005 (2 d.p.) **d** 1.9996 (3 d.p.)

1.3 Converting units

Units are defined so that, for example, every scientist who measures a mass in kilograms uses the same size for the kilogram and gets the same value for the mass. Scientific measurement depends on standard units – most are *Système International* (SI) units.

If you convert between units and round numbers properly it allows quoted measurements to be understood within the scale of the observations.

Multiplication factor	Prefix	Symbol
10^9	giga	G
10^6	mega	M
10^3	kilo	k
10^{-2}	centi	c

10^{-3}	milli	m
10^{-6}	micro	µ
10^{-9}	nano	n

Unit conversions are common. For instance, you could be converting an enthalpy change of 488 889 J mol^{-1} into kJ mol^{-1}. A kilo is 10^3 so you need to divide by this number or move the decimal point three places to the left.

488 889 ÷ 10^3 kJ mol^{-1} = 488.889 kJ mol^{-1}

Converting from mJ mol^{-1} to kJ mol^{-1}, you need to go from 10^3 to 10^{-3}, or move the decimal point six places to the left.

333 mJ mol^{-1} is 0.000 333 kJ mol^{-1}

If you want to convert from 333 mJ mol^{-1} to nJ mol^{-1}, you would have to go from 10^{-9} to 10^{-3}, or move the decimal point six places to the right.

333 mJ mol^{-1} is 333 000 000 nJ mol^{-1}

Practice question

6 Calculate the following unit conversions.

a 300 µm to m

b 5 MJ to mJ

c 10 GW to kW

2 Balancing chemical equations

2.1 Conservation of mass

When new substances are made during chemical reactions, atoms are not created or destroyed – they just become rearranged in new ways. So, there is always the same number of each type of atom before and after the reaction, and the total mass before the reaction is the same as the total mass after the reaction. This is known as the conservation of mass.

You need to be able to use the principle of conservation of mass to write formulae, and balanced chemical equations and half equations.

2.2 Balancing an equation

The equation below shows the correct formulae but it is not balanced.

$H_2 + O_2 \rightarrow H_2O$

While there are two hydrogen atoms on both sides of the equation, there is only one oxygen atom on the right-hand side of the equation against two oxygen atoms on the left-hand side. Therefore, a two must be placed before the H_2O.

$H_2 + O_2 \rightarrow 2H_2O$

Now the oxygen atoms are balanced but the hydrogen atoms are no longer balanced. A two must be placed in front of the H_2.

$2H_2 + O_2 \rightarrow 2H_2O$

The number of hydrogen and oxygen atoms is the same on both sides, so the equation is balanced.

Practice question

1 Balance the following equations.

a $C + O_2 \rightarrow CO$

b $N_2 + H_2 \rightarrow NH_3$

c $C_2H_4 + O_2 \rightarrow H_2O + CO_2$

2.3 Balancing an equation with fractions

To balance the equation below:

$C_2H_6 + O_2 \rightarrow CO_2 + H_2O$

- Place a two before the CO_2 to balance the carbon atoms.
- Place a three in front of the H_2O to balance the hydrogen atoms.

  $C_2H_6 + O_2 \rightarrow 2CO_2 + 3H_2O$

There are now four oxygen atoms in the carbon dioxide molecules plus three oxygen atoms in the water molecules, giving a total of seven oxygen atoms on the product side.

- To balance the equation, place three and a half in front of the O_2.

  $C_2H_6 + 3\frac{1}{2}O_2 \rightarrow 2CO_2 + 3H_2O$

- Finally, multiply the equation by 2 to get whole numbers.

  $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$

Practice question

2 Balance the equations below.

a $C_6H_{14} + O_2 \rightarrow CO_2 + H_2O$

b $NH_2CH_2COOH + O_2 \rightarrow CO_2 + H_2O + N_2$

2.4 Balancing an equation with brackets

$Ca(OH)_2 + HCl \rightarrow CaCl_2 + H_2O$

Here the brackets around the hydroxide (OH^-) group show that the $Ca(OH)_2$ unit contains one calcium atom, two oxygen atoms, and two hydrogen atoms.

To balance the equation, place a two before the HCl and another before the H_2O.

$Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$

Practice question

3 Balance the equations below.

a $Mg(OH)_2 + HNO_3 \rightarrow Mg(NO_3)_2 + H_2O$

b $Fe(NO_3)_2 + Na_3PO_4 \rightarrow Fe_3(PO_4)_2 + NaNO_3$

3 Rearranging equations and calculating concentrations

3.1 Rearranging equations

In chemistry, you sometimes need to rearrange an equation to find the desired values.

For example, you may know the amount of a substance (*n*) and the mass of it you have (*m*), and need to find its molar mass (*M*).

The amount of substance (*n*) is equal to the mass you have (*m*) divided by the molar mass (*M*):

$$n = \frac{m}{M}$$

You need to rearrange the equation to make the molar mass (*M*) the subject.

Multiply both sides by the molar mass (*M*):

$M \times n = m$

Then divide both sides by the amount of substance (*n*):

$$m = \frac{M}{n}$$

Practice questions

1 Rearrange the equation $c = \frac{n}{V}$ to make:

a *n* the subject of the equation

b *V* the subject of the equation.

2 Rearrange the equation *PV* = *nRT* to make:

a *n* the subject of the equation

b *T* the subject of the equation.

3.2 Calculating concentration

The concentration of a solution (a solute dissolved in a solvent) is a way of saying how much solute, in moles, is dissolved in 1 dm^3 or 1 litre of solution.

Concentration is usually measured using units of mol dm^{-3}. (It can also be measured in g dm^3.)

The concentration of the amount of substance dissolved in a given volume of a solution is given by the equation:

$$c = \frac{n}{V}$$

where *n* is the amount of substance in moles, *c* is the concentration, and *V* is the volume in dm^3.

The equation can be rearranged to calculate:

- the amount of substance *n*, in moles, from a known volume and concentration of solution
- the volume *V* of a solution from a known amount of substance, in moles, and the concentration of the solution.

Practice questions

3 Calculate the concentration, in mol dm^{-3}, of a solution formed when 0.2 moles of a solute is dissolved in 50 cm^3 of solution.

4 Calculate the concentration, in mol dm^{-3}, of a solution formed when 0.05 moles of a solute is dissolved in 2.0 dm^3 of solution.

5 Calculate the number of moles of NaOH in an aqueous solution of 36 cm^3 of 0.1 mol dm^{-3}.

4 Molar calculations

4.1 Calculating masses and gas volumes

The balanced equation for a reaction shows how many moles of each reactant and product are involved in a chemical reaction.

If the amount, in moles, of one of the reactants or products is known, the number of moles of any other reactants or products can be calculated.

The number of moles (*n*), the mass of the substance (*m*), and the molar mass (*M*) are linked by:

$$n = \frac{m}{M}$$

Note: The molar mass of a substance is the mass per mole of the substance. For $CaCO_3$, for example, the atomic mass of calcium is 40.1, carbon is 12, and oxygen is 16. So the molar mass of $CaCO_3$ is:

40.1 + 12 + (16 × 3) = 100.1. The units are g mol^{-1}.

Look at this worked example. A student heated 2.50 g of calcium carbonate, which decomposed as shown in the equation:

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

The molar mass of calcium carbonate is 100.1 g mol^{-1}.

a Calculate the amount, in moles, of calcium carbonate that decomposes.

$$n = \frac{m}{M} = 2.50/100.1 = 0.025 \text{ mol}$$

b Calculate the amount, in moles, of carbon dioxide that forms.

From the balanced equation, the number of moles of calcium carbonate = number of moles of carbon dioxide = 0.025 mol

Practice questions

1 In a reaction, 0.486 g of magnesium was added to oxygen to produce magnesium oxide.

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

a Calculate the amount, in moles, of magnesium that reacted.

b Calculate the amount, in moles, of magnesium oxide made.

c Calculate the mass, in grams, of magnesium oxide made.

2 Oscar heated 4.25 g of sodium nitrate. The equation for the decomposition of sodium nitrate is:

$2NaNO_3(s) \rightarrow 2NaNO_2(s) + O_2(g)$

a Calculate the amount, in moles, of sodium nitrate that reacted.

b Calculate the amount, in moles, of oxygen made.

3 0.500 kg of magnesium carbonate decomposes on heating to form magnesium oxide and carbon dioxide. Give your answers to 3 significant figures.

$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$

a Calculate the amount, in moles, of magnesium carbonate used.

b Calculate the amount, in moles, of carbon dioxide produced.

5 Percentage yields

5.1 Calculating percentage yield

Chemists often find that an experiment makes a smaller amount of product than expected. They can predict the amount of product made in a reaction by calculating the percentage yield.

The **yield** of a reaction is the actual mass of product obtained. The **percentage yield** can be calculated: $\%\,yield = \frac{actual\ yield}{theoretical\ yield} \times 100$

Reactions rarely produce 100% yield. Reasons for this could be the reaction is reversible, or side reactions are occurring or there are errors in experimental procedures.

Look at this worked example. A student added ethanol to propanoic acid to make the ester, ethyl propanoate, and water.

$C_2H_5OH + C_2H_5COOH \rightarrow C_2H_5COOC_2H_5 + H_2O$

The experiment has a theoretical yield of 5.00 g.

The actual yield is 4.50 g.

The molar mass of $C_2H_5COOC_2H_5$ = 102.0 g mol^{-1}

Calculate the percentage yield of the reaction.

Actual amount of ethyl propanoate: = 4.5/102 = 0.0441 mol

Theoretical amount of ethyl propanoate: = 5.0/102 = 0.0490 mol

percentage yield = (0.0441/0.0490) × 100% = 90%

Atom economy is a measure of the **amount** of **starting materials** that end up as **useful products**. It is important for **sustainable development** and for **economic reasons** to use reactions with **high atom economy**. The percentage atom economy is calculated using the following equation:

$$\frac{\text{Relative formula mass of desired product from equation}}{\text{Sum of relative formula mass of all reactants from equation}} \times 100$$

Practice questions

1 Calculate the percentage yield of a reaction with a theoretical yield of 4.75 moles of product and an actual yield of 3.19 moles of product. Give your answer to 3 significant figures.

2 Calculate the percentage yield of a reaction with a theoretical yield of 12.00 moles of product and an actual yield of 6.25 moles of product. Give your answer to 3 significant figures.

5.3 Calculating percentage error in apparatus

The percentage error of a measurement is calculated from the maximum error for the piece of apparatus being used and the value measured:

$$\text{percentage error} = \frac{\text{maximum error}}{\text{measured value}} \times 100\%$$

Look at this worked example. In an experiment to measure temperature changes, an excess of zinc powder was added to 50 cm^3 of copper(II) sulfate solution to produce zinc sulfate and copper.

$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

The measuring cylinder used to measure the copper(II) sulfate solution has a maximum error of ±2 cm^3.

a Calculate the percentage error.

percentage error = (2/50) × 100% = 4%

b A thermometer has a maximum error of ±0.05 °C.

Calculate the percentage error when the thermometer is used to record a temperature rise of 3.9 °C. Give your answer to 3 significant figures.

percentage error = (2 × 0.05)/3.9 × 100% = 2.56%

(Notice that two measurements of temperature are required to calculate the temperature change so the maximum error is doubled.)

Practice questions

1 A gas syringe has a maximum error of ±0.5 cm^3. Calculate the maximum percentage error when recording these values. Give your answers to 3 significant figures.

a 21.0 cm^3 **b** 43.0 cm^3

2 A thermometer has a maximum error of ±0.5 °C. Calculate the maximum percentage error when recording these temperature rises. Give your answers to 3 significant figures.

a 12.0 °C **b** 37.6 °C

6. Avogadro Constant

The **relative atomic mass (A_r)** is the weighted average of the masses of its isotopes relative to 1/12 of the mass of a carbon-12 atom. The relative atomic masses can be found in the periodic table.

The **relative molecular mass (M_r)** is the sum of the relative atomic masses of the atoms in the numbers shown in the formula.

In a balanced chemical equation, the sum of the relative formula masses of the **reactants equals** the **sum** of the relative formula masses of the **products**.

For example: $2Mg + O_2 \rightarrow 2MgO$

$(2x24) + (2x16) \rightarrow 2 \times (24+16)$

$80 \rightarrow 80$

Chemical amounts are measured in **moles**. The symbol for the unit mole is **mol**.

The mass of **one mole** of a substance **in grams** is numerically **equal** to its **relative formula mass. One mole of a substance contains the same number of the stated particles, atoms, molecules or ions as one mole of any other substance.**

The **number** of atoms, molecules or ions in a mole of a given substance is the **Avogadro constant**. The value of the Avogadro constant is **6.02×10^{23} per mole**.

For example, 1 mole of H_2O has a mass of 18g and contains : 6.02×10^{23} water molecules, 6.02×10^{23} oxygen atoms and 1.204×10^{24} hydrogen atoms (2 x Avogadro constant).

For simple calculation questions, you need to be able to recall, use and rearrange the following equation:

Number of moles = $\frac{\text{mass (g)}}{A_r}$ or $\frac{\text{mass (g)}}{M_r}$

Question:

Can you calculate the number of magnesium and chloride ions in 50g of magnesium chloride?

7. Empirical and Molecular Formula

The **empirical formula** tells you the **simplest ratio** of the various atoms present in a substance. For example, the empirical formula of ethane (which has a molecular formula of C_2H_6) would be CH_3.

The **molecular formula** gives the **total number of atoms of each element** present in a molecule of the substance. If the empirical formula and relative molecular mass is known, the molecular formula can be calculated.

For example, the empirical formula of ribose is CH_2O. The molar mass of this compound was determined to be 150g/mol. What is the molecular formula of ribose?

Step 1: Determine the molar mass of the empirical formula → 12 + (2 x 1) + 16 = 30gmol^{-1}

Step 2: Divide the given molar mass by your answer from step 1 → 150gmol^{-1} / 30gmol^{-1} = 5

Step 3: Multiply your empirical formula by your answer from step 2 → $C_{1x5}H_{2x5}O_{1x5}$ = $\mathbf{C_5H_{10}O_5}$

Question:

A hydrated salt is analysed and has the following percentage composition by mass:
Cr, 19.51%; Cl, 39.96%; H,4.51%; O, 36.02%
Calculate the formula of the compound showing clearly the water of crystallisation.

Appendix 2 Structure and Bonding

a. Ionic

Ionic bonds form between **metals and non-metals**. Ionic bonding involves the **transfer of electrons** in the **outer** shells.

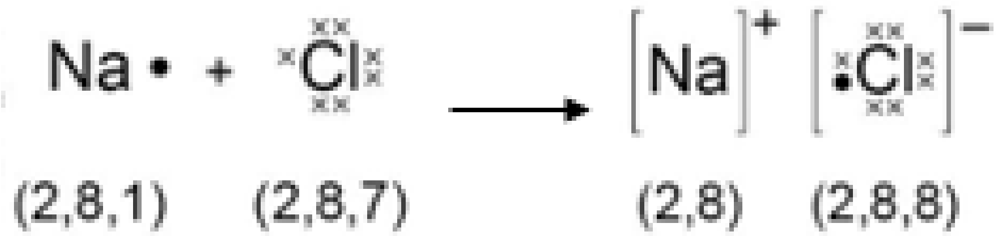
Metals **lose** electrons to become **positively** charged ions.

Non-metals **gain** electrons to become **negatively** charged ions.

Learn these:

Positive Ions		Negative Ions	
Hydrogen	H^+	Fluoride	F^-
Lithium	Li^+	Chloride	Cl^-
Sodium	Na^+	Bromide	Br^-
Potassium	K^+	Iodide	I^-
Magnesium	Mg^{2+}	Oxide	O^{2-}
Calcium	Ca^{2+}	Hydroxide	OH^-
Aluminium	Al^{3+}	Nitrate	NO_3^-
Silver	Ag^+	Sulphate	SO_4^{2-}
Copper	Cu^{2+}	Phosphate	PO_4^{3-}
Ammonium	NH_4^+	Carbonate	CO_3^{2-}
Iron	Fe^{2+} & Fe^{3+}		
These have all lost electrons. They're all metals apart from H^+ and NH_4^+		These have all gained electrons. They're all non-metals.	

The **electrostatic attraction** between the oppositely charged ions is called **ionic bonding**. The electron transfer during the formation of an ionic compound can be represented by a **dot and cross diagram**:

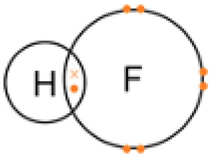


Question

Explain why magnesium oxide has a higher melting point than sodium chloride.

b. Covalent

Covalent bonds form between **non-metals**. Covalent bonding is a **shared pair of electrons** in the **outer** shells. Therefore a single bond is one shared pair of electrons, a double bond is two shared pairs of electrons etc.



Covalent bonding is the **electrostatic attraction** between **shared pair of electrons** and two **positive nuclei.**

Covalent bonds should **not** be regarded as **shared electron pairs in a fixed position**; the electrons are in a **state of constant motion** and are best regarded more as **charge clouds**.

A **dative covalent bond** (or **coordinate bond**) is a **pair of electrons shared between two** atoms, **one of which provides both electrons** to the bond.

A dative covalent bond is represented by a short arrow from the electron providing both electrons to the electron providing neither.

Question

Name the type of bond formed between N and Al in H_3NAlCl_3 and explain how this bond is formed.

c. Metallic

A **metallic bond** is an **attraction** between **positive ions** and a **sea of delocalised electrons**.

Metallic bonds are formed when **atoms lose electrons** and the resulting **electrons** are **attracted** to **all** the **resulting cations**. Metallic bonding happens because the **electrons** are **attracted** to **more than one nucleus** and hence **more stable**. The **electrons** are **delocalised** – they are not attached to any particular atom, but are free to move between the atoms.

Question

Explain why aluminium has a higher melting point than sodium.

d. Properties of Substances

Ionic:

- The **attraction** between opposite ions is **very strong**. **A lot** of kinetic **energy** is **required** to overcome them and the **melting point** and **boiling point** of ionic compounds is **very high**.

- Since ions are held strongly in place by the other ions, they **cannot move or slip over each other** easily and are therefore **hard and brittle**.

- Ionic compounds **contain charged ions** so they are able to move towards charged electrodes and will **therefore conduct electricity**. In the **solid state** the ions are **not free to move** as they are tightly held in place. They **do not conduct** electricity. In the liquid state, the ions are free to move and so can move towards their respective electrodes. Ionic compounds can conduct electricity in the liquid state.

Metallic:

- Metallic bonding is relatively strong so the melting and boiling points of metals are relatively high. Smaller ions, and those with a high charge, attract the electrons more strongly and so have higher melting points than larger ions with a low charge.

- Delocalised electrons are free to move throughout the crystal in a certain direction when a potential difference is applied. Metals therefore conduct electricity in the solid state. The delocalised electron system is still present in the liquid state, so metals can also conduct electricity well in the liquid state.

- Metal cations can be moved around and there will still be delocalised electrons available to hold the cations together. The metal cations can therefore slip over each other fairly easily. As a result, metals tend to be soft, malleable and ductile.

Simple /molecular covalent:

- Melting and boiling points are generally low, since intermolecular forces are weak. Intermolecular forces also decrease rapidly with increasing distance, so there is often little difference in the melting and boiling points.

- There are no ions and no delocalised electrons, so there is little electrical conductivity in either solid or liquid state.

- The intermolecular forces are weak and generally non-directional, so most molecular covalent substances are soft, crumbly and not very strong.

Giant covalent – diamond and silicon dioxide:

- Generally very high melting and boiling points, since strong covalent bonds must be broken before any atoms can be separated.

- There are no ions or delocalised electrons, so there is little electrical conductivity in either solid or liquid state.

- Giant covalent substances are hard, strong and brittle.

Giant covalent – graphite:

- Due to the delocalised electrons, graphite is a very good conductor of electricity.

- Graphite has a much lower density than diamond due to the relatively large distances in between the planes.

- Much softer than diamond since the different planes can slip over each other fairly easily. This results in the widespread use of graphite in pencils and as an industrial lubricant.

Question

In terms of structure and bonding, explain why the boiling point of bromine is different from that of magnesium.

Appendix 3 A Level definitions (3 hours)

Retrieval questions

You need to be confident about the definitions of terms that describe measurements and results in A Level Chemistry.

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many answers as you can. Check and repeat.

Practical science key terms

When is a measurement valid?	when it measures what it is supposed to be measuring
When is a result accurate?	when it is close to the true value
What are precise results?	when repeat measurements are consistent/agree closely with each other
What is repeatability?	how precise repeated measurements are when they are taken by the *same* person, using the *same* equipment, under the *same* conditions
What is reproducibility?	how precise repeated measurements are when they are taken by *different* people, using *different* equipment
What is the uncertainty of a measurement?	the interval within which the true value is expected to lie
Define measurement error	the difference between a measured value and the true value
What type of error is caused by results varying around the true value in an unpredictable way?	random error
What is a systematic error?	a consistent difference between the measured values and true values
What does zero error mean?	a measuring instrument gives a false reading when the true value should be zero

Which variable is changed or selected by the investigator?	independent variable
What is a dependent variable?	a variable that is measured every time the independent variable is changed
Define a fair test	a test in which only the independent variable is allowed to affect the dependent variable
What are control variables?	variables that should be kept constant to avoid them affecting the dependent variable

Atoms, ions, and compounds

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many answers as you can. Check and repeat.

What does an atom consist of?	a nucleus containing protons and neutrons, surrounded by electrons
What are the relative masses of a proton, neutron, and electron?	1, 1, and respectively
What are the relative charges of a proton, neutron, and electron?	+1, 0, and -1 respectively
How do the number of protons and electrons differ in an atom?	they are the same because atoms have neutral charge
How does the number of protons differ between atoms of the same element?	it does not differ – all atoms of the same element have the same number of protons
What force holds an atom nucleus together?	strong nuclear force
What is the proton number / atomic number of an element?	the number of protons in the atom's nucleus of an element
What is the mass number of an element?	number of protons + number of neutrons
What is an isotope?	an atom with the same number of protons but different number of neutrons

What is the equation for relative isotopic mass?	relative isotopic mass = $\dfrac{\text{mass of an isotope}}{\frac{1}{12}^{\text{th}} \text{ mass of 1 atom of } ^{12}\text{C}}$
What is the equation for relative atomic mass (A_r)?	relative atomic mass = $\dfrac{\text{weighted mean mass of 1 atom}}{\frac{1}{12}^{\text{th}} \text{ mass of 1 atom of } ^{12}\text{C}}$
What is the equation for relative molecular mass (M_r)?	relative molecular mass = $\dfrac{\text{average mass of 1 molecule}}{\frac{1}{12}^{\text{th}} \text{ mass of 1 atom of } ^{12}\text{C}}$
What is an ion?	an atom or group of atoms with a charge (a different number of electrons to protons)
Define the term cation	a positive ion (atom with fewer electrons than protons)
Define the term anion	a negative ion (atom with more electrons than protons)
What is the function of a mass spectrometer?	it accurately determines the mass and abundance of separate atoms or molecules, to help us identify them
What is a mass spectrum?	the output from a mass spectrometer that shows the different isotopes that make up an element
What is a binary compound?	a compound which contains only two elements

Some answers to the Appendix 1

Maths skills

1 Core mathematics

Practice questions

1 **a** 1.413×10^3 °C **b** 1.0×10^{-7} m
c 1.806×10^{21} atoms

2 **a** 0.000 0055 **b** 290
c 11150 **d** 0.001 412
e 72

3 **a** 36.9 **b** 260
c 0.043 **d** 8 000 000

4 Number of molecules = 0.5 moles × $6.022 \times 10^{23} = 3.011 \times 10^{23} = 3.01 \times 10^{23}$

5 **a** 4.8 **b** 0.54
c 1.01 **d** 2.000

6 **a** 0.0003 m **b** 5×10^9 mJ
c 1×10^7 kW

2 Balancing chemical equations

Practice questions

1 **a** $2C + O_2 \rightarrow 2CO$ **b** $N_2 + 3H_2 \rightarrow 2NH_3$
c $C_2H_4 + 3O_2 \rightarrow 2H_2O + 2CO_2$

2 **a** $C_6H_{14} + 9\frac{1}{2}O_2 \rightarrow 6CO_2 + 7H_2O$ **or** $2C_6H_{14} + 19O_2 \rightarrow 12CO_2 + 14H_2O$
b $2NH_2CH_2COOH + 4\frac{1}{2}O_2 \rightarrow 4CO_2 + 5H_2O + N_2$
or $4NH_2CH_2COOH + 9O_2 \rightarrow 8CO_2 + 10H_2O + 2N_2$

3 **a** $Mg(OH)_2 + 2HNO_3 \rightarrow Mg(NO_3)_2 + 2H_2O$
b $3Fe(NO_3)_2 + 2Na_3PO_4 \rightarrow Fe_3(PO_4)_2 + 6NaNO_3$

3 Rearranging equations and calculating concentrations

Practice questions

1 **a** $n = cv$ **b** $v = \frac{n}{c}$

2 **a** $n = \frac{PV}{RT}$ **b** $T = \frac{PV}{nR}$

3 $\frac{0.2}{0.050} = 4.0 \text{ mol dm}^{-3}$

4 $\frac{0.05}{2} = 0.025 \text{ mol dm}^{-3}$

5 $\frac{36}{1000} \times 0.1 = 3.6 \times 10^{-3} \text{ mol}$

4 Molar calculations

Practice questions

1 **a** $\frac{0.486}{24.3} = 0.02$ mol **b** 0.02 mol

c 0.02 × 40.3 = 0.806 g

2 **a** $\frac{4.25}{85} = 0.05$ mol **b** $\frac{0.05}{2} = 0.025$ mol

3 **a** $\frac{500}{84.3} = 5.93$ mol **b** 5.93 mol

5 Percentage yields and percentage errors

Practice questions

1 3.19/4.75 × 100 = 67.2%

2 6.25/12.00 × 100 = 52.1%

3 **a** 0.5/21 × 100 = 2.38% **b** 0.5/43 × 100 = 1.16%

4 **a** 0.5 × (2/12) × 100 = 8.33% **b** 0.5 × (2/37.6) × 100 = 2.66%