Kirkcaldy High School

Chemistry

National 5

Unit 1 – Chemical Changes and Structure

NOTES

Course Overview

Contents

The National 5 Chemistry Course is split into three units. *Italic* shows the contents of this notes booklet.

Unit 1 – Chemical Changes in Structure

Unit 2 – Nature's Chemistry

(a) Homologous Series
(b) Everyday consumer products
(c) Energy from Fuels

Unit 3 – Chemistry in Society

(a) Metals
(b) Plastics
(c) Fertilisers
(d) Nuclear chemistry
(e) Chemical analysis

Assessment

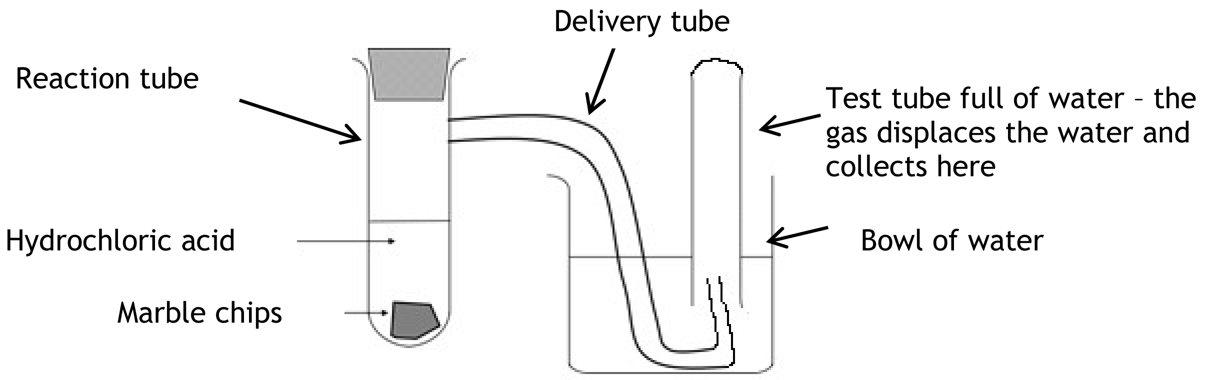
- There is a **final exam** for National 5 Chemistry. It lasts for **2 h 30 mins** and contains **100 marks**. There are:
  - 25 marks of multiple-choice questions
  - 75 marks of written answer questions
- There is an **Assignment** (essay on a Chemistry topic) to write for National 5 Chemistry.
  - The Assignment is **externally marked** by an SQA marker (NOT your teacher)
  - The Assignment contains **20 marks,** but these are scaled to 25 marks such that the assignment is worth one quarter of the total marks.
  - You have unlimited **time to research** and gather data for your Assignment.
  - You can **1 h 30 mins to write** the Assignment.

(a) Rates of Reaction

Tests for gases

- The test for hydrogen gas - burns with a squeaky pop
- The test for oxygen gas - relights a glowing splint
- The test for carbon dioxide gas - turns limewater cloudy

Method Collecting a gas from a reaction



Measuring Rate of Reaction

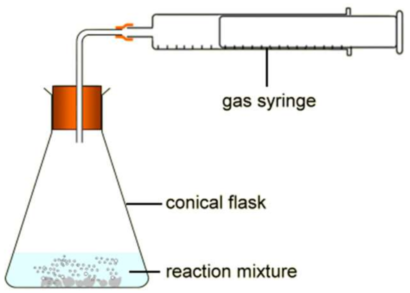
The speed of a reaction is referred to as the 'rate of a reaction'

The rate of a reaction can be measured by recording the:

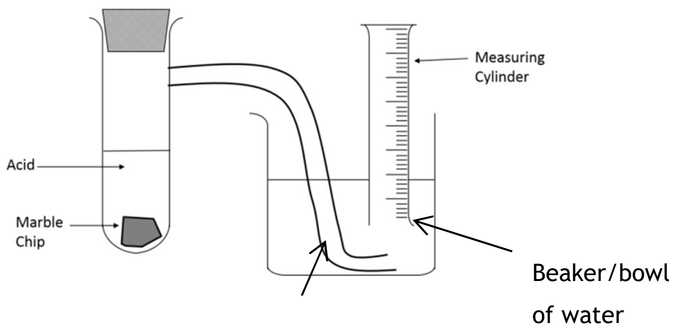
(a) Change in the **mass** of the reactants or products over a given time OR

(b) Change in the **volume** of the reactants or products over a given time OR

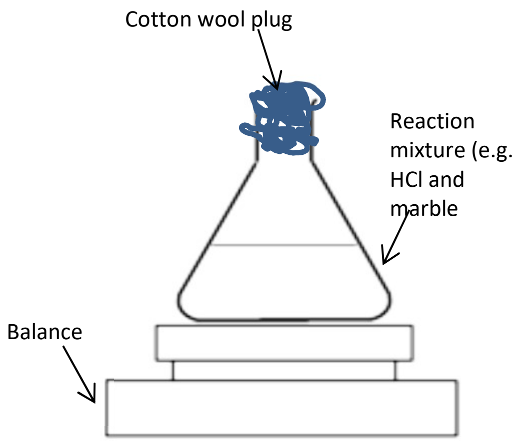
Measuring the change in volume:



OR



Measuring the change in mass:



The equation for finding the average rate of a reaction over a given time

$$\textbf{Average Rate} = \frac{\textbf{change in quantity}}{\textbf{change in time}} = \frac{\Delta\textbf{quantity}}{\Delta\textbf{time}}$$

(Measurable quantity = mass, volume or concentration)

You must be able to calculate average rate from numbers in text, data in a table and from a graph (see examples in notes and sheet of graphs)

Example units:

$$\frac{\text{change in volume}}{\text{change in time}} = cm^3/s \text{ or } cm^3s^{-1}$$

$$\frac{\text{change in mass}}{\text{change in time}} = g/s \text{ or } gs^{-1}$$

Example calculation:

Time /s	**0**	**10**	**20**	**40**	**60**	**80**
Volume of gas /cm^3	0	17	26	71	106	135

Calculate the average rate of reaction between 20 seconds and 60 seconds.

Average rate =

$$\frac{\textbf{change in volume}}{\textbf{change in time}}$$

$$\text{Average rate} = \frac{106\text{-}26}{60\text{-}20}$$

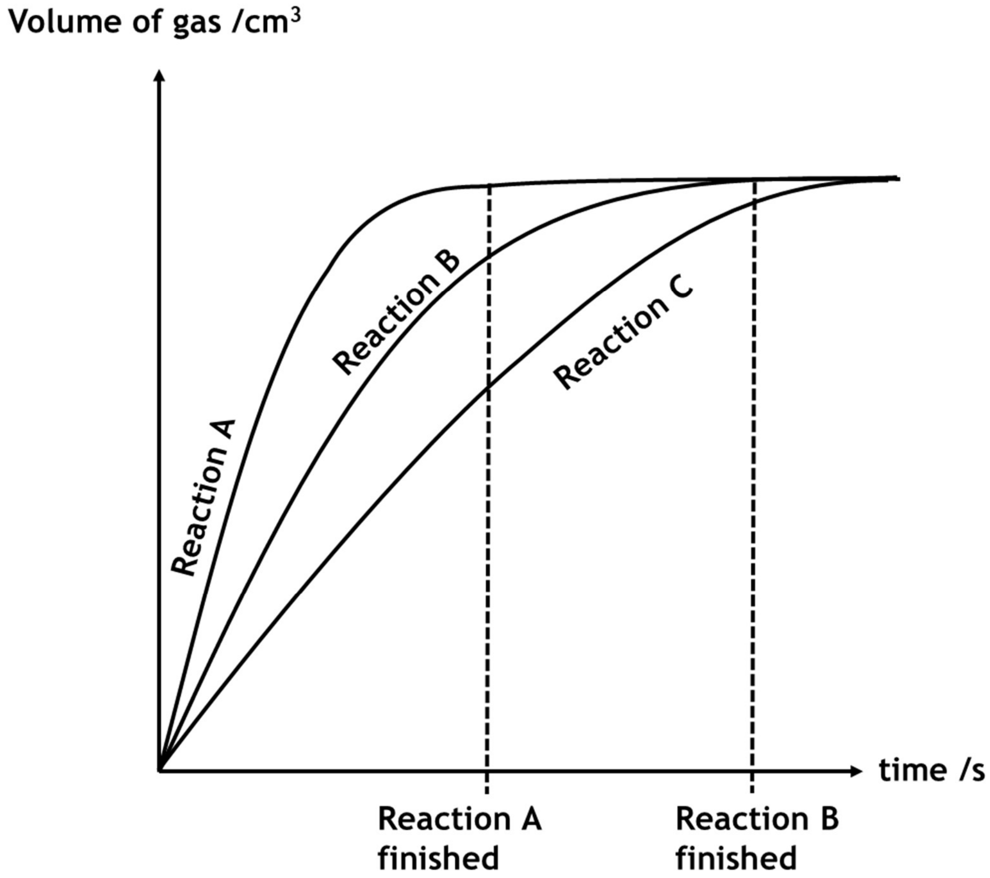
$$\text{Average rate} = \frac{80}{40}$$

Average rate = 2 cm^3s^{-1}

Average rates of reaction over various time intervals can show that as the reaction progresses the rate of reaction decreases (the slope/gradient of a line gives an indication of the speed) This is because the concentration of the reactants is decreasing.

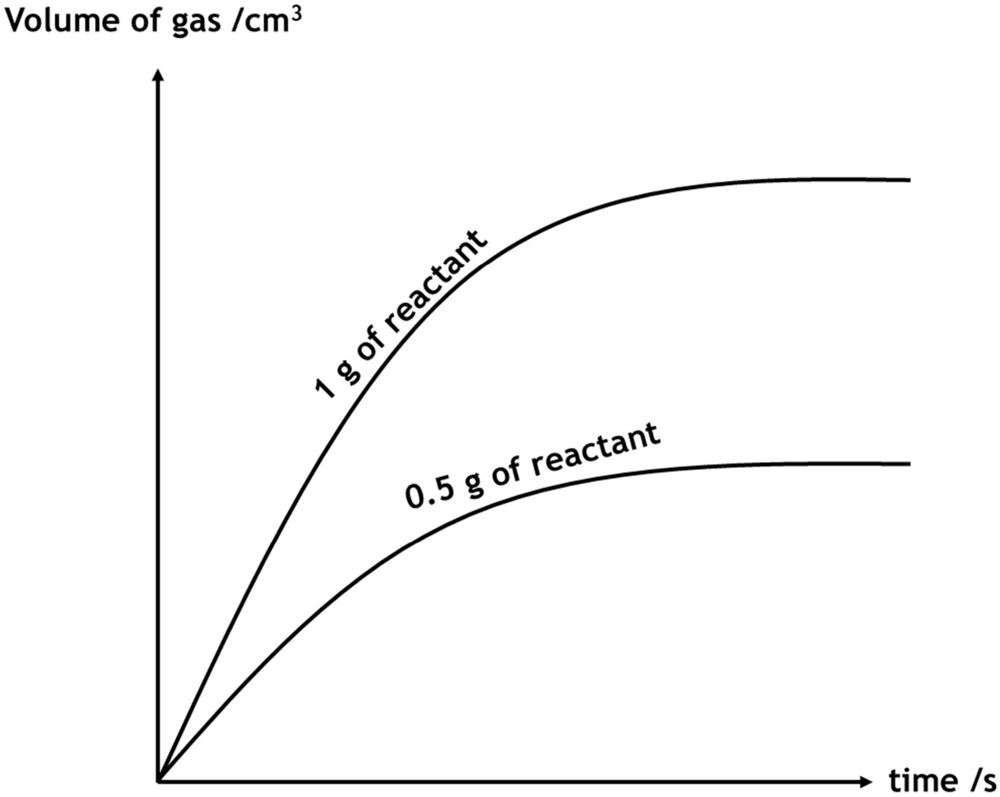
Reaction progress graphs

Volume vs Time graph:

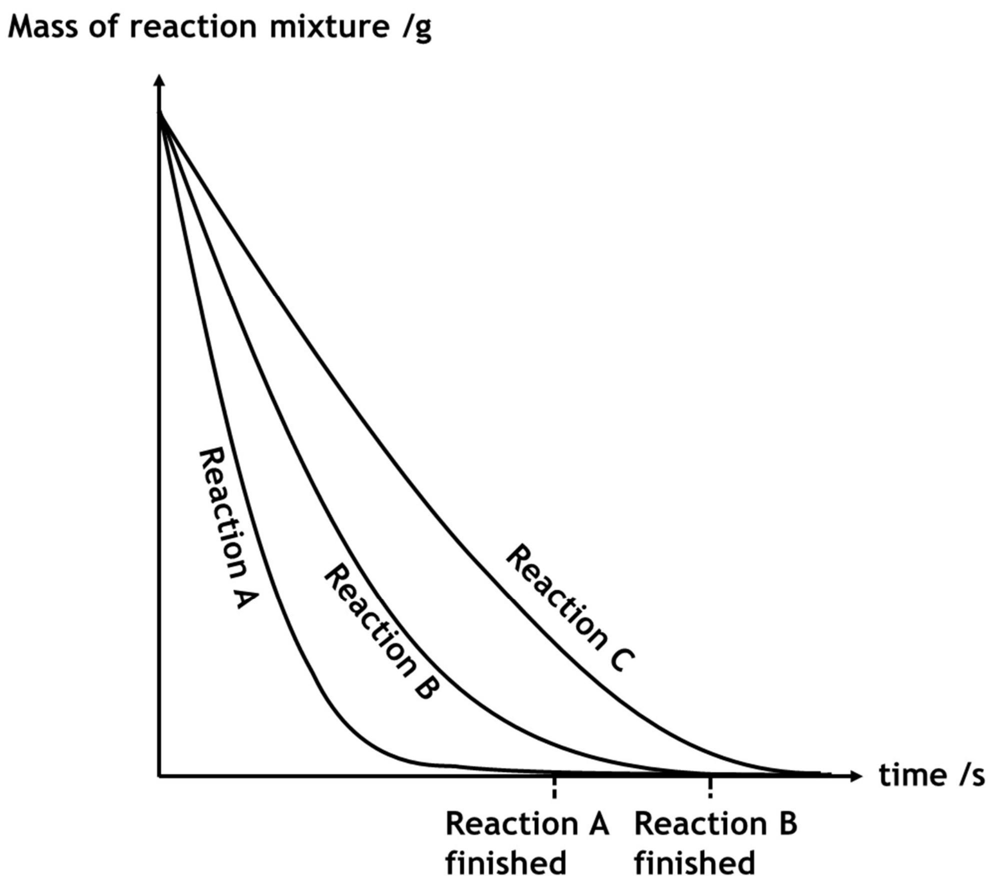


Reaction A:	**Steep gradient**	**faster reaction**		**smaller particle size**
			OR	higher temperature
			OR	higher concentration
			OR	with catalyst
Reaction C:	Less steep gradient	Slower reaction		larger particle size
			OR	lower temperature
			OR	lower concentration
			OR	without catalyst

Graph to show the outcome if the mass of the reactants changes:



Mass vs Time graph:



Reaction A:	**Steep gradient**	**faster reaction**		**smaller particle size**
			OR	higher temperature
			OR	higher concentration
			OR	with catalyst
Reaction C:	Less steep gradient	Slower reaction		larger particle size
			OR	lower temperature
			OR	lower concentration
			OR	without catalyst

Changing the Rate

For a chemical reaction to take place the reacting particles must **COLLIDE** with enough **ENERGY** to combine and **form new products** (***collision theory***). Changing one of the factors below causes either more collisions, increases the energy of the collision or both.

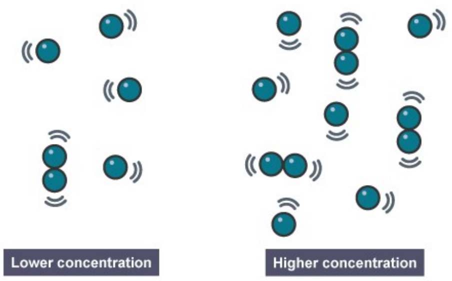
The **factors** that affect the rate of a reaction are:

- Temperature of the reactants
- Particle size of the reactants
- Concentration of the reactants
- Use of a catalyst

Temperature: If the temperature is increased the particles have more energy and move more quickly. This increases the rate of reaction because particles collide more often and with more energy.

- Higher the temperature = Faster the reaction
- Lower the temperature = Slower the reaction

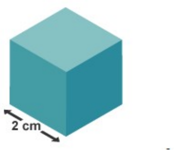
Concentration: If the concentration of the reactants is increased there are more reactant particles moving around in the same space. This increases the rate of reaction because there will be more collisions.



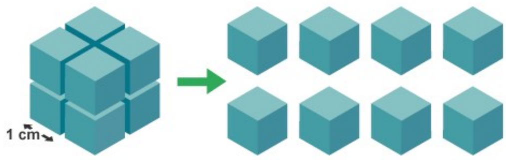
- Higher the concentration = Faster the reaction
- Lower the concentration = Slower the reaction

Particle Size: If the particle size is decreased (*i.e.* lumps into powder) there is an increase the surface area. This increases the rate of reaction because a greater surface area means a higher chance of collisions occurring.

- Smaller particle size = Faster the reaction
- Larger particle size = Slower the reaction



Larger particle size, smaller surface area



Smaller particle size, larger surface area

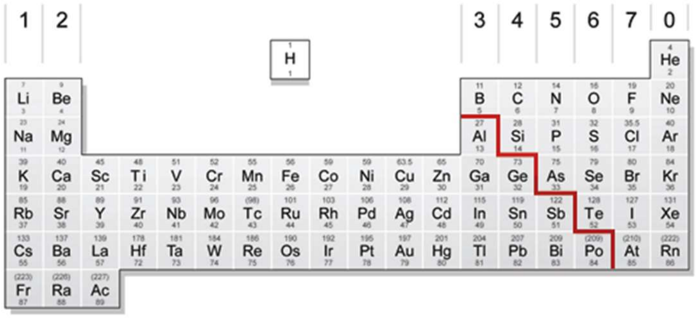
Catalysts: are substances used to speed up the rate of the reaction but are not used up in the reaction (i.e. can be recovered and re-used at the end of the reaction). Example: platinum used in car exhausts. Advantages: save time, heating, and money

(b) Atomic structure and bonding related to properties of materials

Periodic Table and atoms

Periodic Table

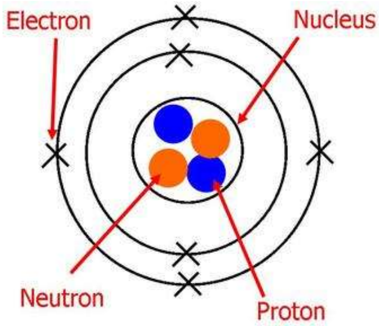
- Elements can be identified by their name or by a symbol.
- Elements are arranged in the periodic table.
- The periodic table is made up of periods (rows) and groups (columns).
- Elements in the same group of the periodic table have similar chemical properties due to their electron configurations (*e.g.* the first group all react violently with water and all have 1 electron in their outer-shell of electrons).
- **Metals** are on the **left hand side** of the zig-zag in the periodic table and **non-metals** are on the **right hand side** of the zig-zag in the periodic table.
- Elements are made up of only one type of atom
- Elements are arranged in order of increasing atomic number



Atoms

Atoms are made up of 3 sub-atomic particles:

- Protons
- Neutrons
- Electrons



Particle	Charge	Mass	Location
<u>P</u>rotons	+1 (**<u>p</u>**ositive)	1	Inside nucleus
<u>Neu</u>trons	0 (**<u>neu</u>**tral)	1	Inside nucleus
<u>E</u>lectrons	-1 (n**<u>e</u>**gative)	0	Outside nucleus

The centre of an atom is called the nucleus. The nucleus is made up of protons and neutrons and contains all the mass of the atom. The electrons are in shells (or energy levels) outside of the nucleus.

Atoms are neutral because the number of positive protons is equal to the number of negative electrons

Atomic number = the number of protons (Data Book page 4)

Mass number = the number of protons + the number of neutrons (Data book page 7)

Number of neutrons = Mass number – Atomic number

Nuclide notation

In this example, there are 2 protons, 2 neutrons (4-2) and 2 electrons (since it is neutral protons = electrons)

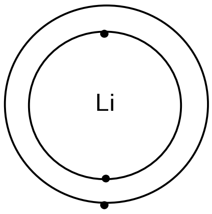
Electron Arrangements

Electrons are arranged in shells around a nucleus. The first shell can contain a maximum of 2 electrons and the next shells contain a maximum of 8

The electron configuration shows how the electrons are arranged in the shells (page 6 db)

e.g. Na: 2,8,1; Ca: 2,8,8,2; F: 2,7

An electron arrangement diagram shows electrons in the shells: Example: Lithium; 2,1



Nobel gases, group 8 – have full outer-shells of electrons making them unreactive

Ions

Ions are Charged particles with different numbers of protons and electrons. Electrons are lost or gained to achieve a stable electron configuration

- Metals form positive ions, since they lose electrons
- Non-metals form negative ions, since they gain electrons

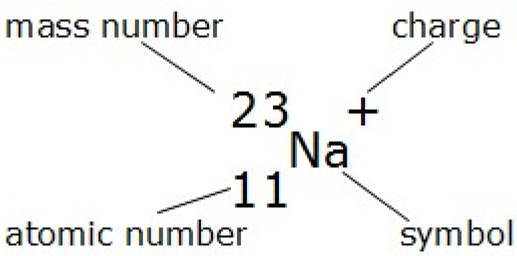
The charge on an ion = you must give a number and a sign, the sign is **+** for metal, **–** for non-metal, the number is the valency number

Valency:

Group number	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**
Valency	1	2	3	4	3	2	1	0

e.g. Ca^{2+}, F^{-}, Na^{+}, O^{2-}, Al^{3+}

Nuclide notation for ions:



This ion has

- 11 protons
- 12 neutrons
- 10 electrons (because 1+ charge means loss of 1 electron)

Isotopes

Isotopes are atoms with the same atomic number but different mass numbers. This means they have the same number of protons but a different number of neutrons. *e.g.* $^{35}_{17}Cl$ and $^{37}_{17}Cl$.

Relative Atomic Mass (RAM) = the average mass of the isotopes present accounting for their relative proportions

$$\text{Relative Atomic Mass} = \frac{(\text{Mass of isotope 1 x \% isotope 1}) + (\text{Mass of isotope 2 x \% isotope 2})}{100}$$

Since Relative Atomic Mass is a weighted average of the isotopes, the isotope closest in value to the Relative Atomic Mass is the one with the highest percentage. *e.g.* for $^{35}_{17}Cl$ and $^{37}_{17}Cl$, the Relative Atomic Mass is 35.5. This tells us there is a higher percentage of the $^{35}_{17}Cl$ isotope because 35 is closer to 35.5 than 37 is.

Compounds

A compound is two or more elements chemically bonded together.
Compounds are named after the elements present in them.

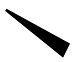
Covalent Bonding

The octet rule of thumb = atoms want to have **full electron arrangements** like the noble gases (group 8)
Molecule = two or more atoms bonded together

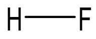
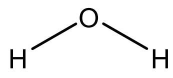
Shapes of molecules

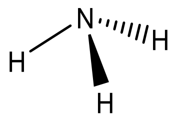
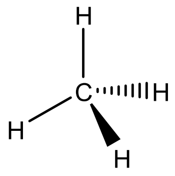
- Linear *e.g.* HF
- Bent *e.g.* H_2O
- Trigonal pyramidal *e.g.* NH_3
- Tetrahedral *e.g.* CH_4

Drawing the shapes of molecules use

-  to show a bond going away from you out of the plane of the paper
-  to show a bond coming towards you out of the plane of the paper
- to show a bond in the plane of the paper

e.g.

Linear:  Bent: 

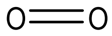
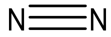
Trigonal pyramidal:  Tetrahedral: 

Diatomic elements

These exist in nature as two atoms covalently bonded; diatomic = **two** atoms [you can also get diatomic compounds which are two atoms chemically bonded together but remember a compound is made up of 2 or more different elements]

7 diatomic elements = Hydrogen (H_2), Nitrogen (N_2), Oxygen (O_2), Fluorine (F_2), Chlorine (Cl_2), Bromine (Br_2), Iodine (I_2) [HON halogen or remember N – I make a 7 on the periodic table]

More than 1 bond can form between atoms leading to double and triple bonds.

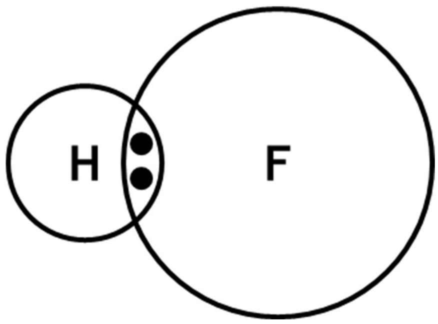
e.g.  

Covalent Bonds

A covalent bond is a **sharing of a pair of electrons** between two **non-metal** atoms. The attraction between a shared pair of negative electrons and the two positive nuclei of the atoms sharing the electrons holds the bond together.

Drawing covalent molecules you draw the outer shell of electrons and show the overlap

e.g.



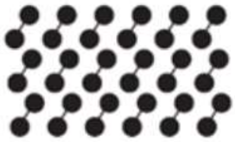
Remember electrons are shared to complete the outer shell of electrons (the first shell can have a maximum of 2 electrons, the second and third shell can have a maximum of 8 electrons)

Valency tells us how many bonds atoms need to form in order to become stable and is linked to an elements outer electrons and therefore group in the periodic table.

- For groups 1-4 the valency is equal to the group number
- For groups 5-8 the valency is [8-group number]

Covalent substances can either form **discrete covalent molecules** or **giant covalent networks**

e.g.



Discrete covalent molecule Giant covalent Network

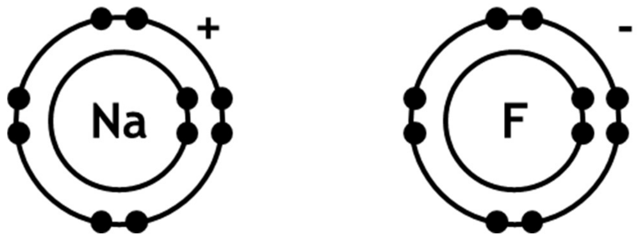
Discrete covalent molecules form when a small number of atoms combine in fixed numbers.
Giant covalent networks form when the number of atoms joining together is huge and the number of atoms is not fixed.
Three covalent network elements: **boron, carbon** and **silicon** (triangle in periodic table)
Covalent network compounds: **silicon dioxide** and **silicon carbide** [graphite and diamond are covalent networks that are made up of just carbon]

Ionic Compounds

Ionic bonding occurs between **metals** and **non-metals**. To complete the outer shell of electrons and become stable there is a transfer of electrons. The resulting atoms are charged and are called ions. Ionic bonds are held together through the electrostatic forces of attraction between positive and negative ions.
Ions are charged particles formed from either the loss or gain of electrons.

- If electrons are lost, a positive ion is formed (**Metals** form **positive** ions)
- If electrons are gained, a negative ion is formed (**Non-metals** form **negative** ions)

The charges on the ions are linked to the number of electrons in the outer shell. For metals it is easier for them to lose the outer shell electrons. Since the group number indicates how many electrons there are in the outer shell; group 1 = 1+ ions (loss of 1 electron), group 2 = 2+ ions (loss of 2 electrons), group 3 = 3+ ions (loss of 3 electrons). For non-metals it is easier for them to gain electrons therefore the charge is [8-group number] i.e. the number of electrons needed to fill the outer shell up to 8 electrons. Group 6 = 2- ions (gain of 2 electron), group 7 = 1- ions (gain of 1 electron).
Remember when asked for the number of electrons an ion has you MUST consider the charge.
e.g. F^- is [F + 1 electron] therefore it has 10 electrons in total (NOT 9 like a F atom) F = 2,7; F^- = 2,8
e.g. Mg^{2+} is [Mg – 2 electrons] therefore it has 10 electrons in total (NOT 12 like a Mg atom) Mg = 2,8,2; Mg^{2+} = 2,8
e.g. NaCl



Sodium has lost 1 electron to become Na^+

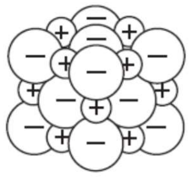
- Na^+ has 10 electrons
  - since Na atoms have 11 electrons but it has lost 1 electron to become an ion;2,8,1 → 2,8

Chlorine has gained 1 electron to become Cl^-

- Cl^- has 18 electrons
  - since Cl atoms have 17 electrons but it has gained 1 electron to become an ion; 2,8,7 → 2,8,8

Ionic Lattice

An **ionic lattice** is the structure that is formed from oppositely charged ions. The ions are charged all the way around and attract oppositely charged ions in all directions forming a regular arrangement of ions.



Ionic compounds are **neutral** overall therefore the charges on the positive metal ions and negative non-metal ions must cancel (i.e. equal 0). The **chemical formula** is generated from this.
e.g.
An ionic compound is formed between calcium and chlorine
Step 1: *work out what charge a calcium ion and a chlorine ion have by considering their outer shell electrons*
Calcium (2,8,8,2) has 2 electrons in the outer shell it will lose these 2 electrons to form Ca^{2+}
Chlorine (2,8,7) has 7 electrons in the outer shell it is easier to gain 1 electron than lose 7 therefore it will gain an electron to complete its outer shell and become Cl^-
Step 2: *balance the charges*
Calcium has a 2+ charge and chlorine has a 1- charge if we had 1 calcium and 1 chlorine overall this would be a positive compound (2+ + (1-) = 1+) but we need it to be neutral therefore we need another 1- to balance overall and become neutral.

$$Ca^{2+}\ Cl^-\ Cl^- = 1 \times Ca^{2+} \text{ and } 2 \times Cl^-$$

Step 3: *write this as a chemical formula*

$CaCl_2$ (this subscript 2 indicates 2 chlorine ions)

You must be able to work backwards if given a formula.

e.g.

X_2O what is the charge on element X?

Step 1: *charge on the known element*

Oxygen ion → O^{2-}

Step 2: *balancing the charge*

We know we have a 2- therefore we need 2+ to balance (i.e. = 0 overall)

Step 3: *consider how many atoms this charge is spread over*

X_2 therefore 2 atoms of X to get 2+ charge overall each X must have a 1+ charge

The **ionic formula** of a compound is the same as the chemical formula but it also includes the charges on the ions.

e.g.

Calcium chloride

Chemical formula = $CaCl_2$ Ionic formula = $\mathbf{Ca^{2+}Cl^{-}_2}$

Properties

Property	Ionic Lattice	Covalent Network	Covalent Molecular
Melting point and boiling point	High	Very High	Low
Conducts electricity	When molten or in solution	No (except graphite)	No
State at room temperature	Solid	Solid	Liquid or gas (or low melting solid)

Ionic compounds only conduct when they are in solution (liquid) or molten. This is because the lattice structure breaks down and the ions (charged particles) are free to move. They also have high melting and boiling points because strong ionic bonds must be broken in order to break down the lattice.

Covalent molecular substances do not tend to conduct electricity. They also have low melting and boiling points because the intermolecular bonds between the covalent molecules (forces of attraction) are weak.

Covalent network substances do not tend to conduct electricity. They also have very high melting and boiling points because strong covalent bonds between all the atoms must be broken.

Measuring melting point or boiling point can be used to indicate the type of bonding and measuring the electrical conductivity can be used to confirm the type of bonding.

(c) Formulae and reacting quantities

(i) Chemical Formulae

•**Naming compounds:**

Two element compounds: the names of the elements are written from left to right as they appear in the periodic table but the ending on the second element is changed to –ide

e.g.

HCl Hydrogen Chloride

MgO Magnesium Oxide

Three element compounds (and one of these is oxygen) the ending becomes –ate.

e.g.

Copper Sulphate ($CuSO_4$) - this contains copper sulphur and oxygen.

Chemical Formula – Prefixes

Prefixes come before a word in a chemical name to tell you how many of that element is present

Prefix = Number of atoms

Mono	**1**
Di	2
Tri	3
Tetra	4
Penta	5
Hexa	6

e.g.

NH_3 nitrogen trihydride

CCl_4 carbon tetrachloride

Chemical Formula – Valencies

Writing chemical formula – "Valency Swap and Drop"

Group number	1	2	3	4	5	6	7	8
Valency	1	2	3	4	3	2	1	0

e.g.

Silicon Oxide

Symbols		**Si**	**O**
Valencies		4	2
		Si	O
Valencies swapped and written subscript		Si_2	O_4
Simplify (covalent networks and ionic compounds only)	÷ 2	Si	O_2
Formula		**SiO_2**	

Compounds containing transition metals – the roman numeral in the brackets is the valency number of the transition metal

Roman Numeral	Valency
I	1
II	2
III	3

Roman Numeral	Valency
IV	4
V	5
VI	6

e.g.

Cobalt (II) Chloride

Symbols	**Co Cl**
Valencies	2 1 Co Cl
Valencies swapped and written subscript	Co Cl_2
Formula	**$CoCl_2$**

Compounds containing group ions – the charge number is the valency number of the group ion (see page 8 databook for group ions)

e.g.

Ion name	Formula	Charge	Valency
Ammonium	NH_4^+	1+	1
Carbonate	CO_3^{2-}	2-	2

Ion name	Formula	Charge	Valency
Hydroxide	OH^-	-1	1
Phosphate	PO_4^{3}	-3	3

e.g.

Calcium Nitrate

Symbols	**Ca**	**NO_3**	
Valencies	2	1	
	Ca	NO_3	
Valencies swapped and written subscript	Ca	$(NO_3)_2$	when a group ion is multiplied it must be put in brackets
Formula	**$Ca(NO_3)_2$**		

The chemical formula of a covalent molecular substance gives the number of atoms present in the molecule. The formula of a covalent network or ionic compounds gives the simplest ratio of atoms/ions in the substance.

Calculations involving the mole and balanced equations

Word equations

Tell you what is used in the reaction and what is made in the reaction

e.g.

Carbon + Oxygen → Carbon Dioxide

Formula equations show the chemical formula

$$C + O_2 \rightarrow CO_2$$

State symbol	Meaning
(s)	Solid
(l)	Liquid
(g)	Gas
(aq)	Aqueous – dissolved in water

State symbols can be used after the name or formula to show the chemical state of the substances.

Balancing Equations

To balance an equation, we must adjust the number of formula units of some of the substances until we get an equal number of each type of atom on both sides of the equation.
Balancing involves writing numbers in front of formulae – NEVER change small numbers within chemical formula
Deal with 1 element at a time
The numbers in front of each reactant tell us the ratio that is required
e.g.

$\mathbf{H_2}$	**+**	$\mathbf{O_2}$	$\rightarrow$	$\mathbf{H_2O}$
H x 2				H x 1
O x 2				O x 1
H_2	+	O_2	$\rightarrow$	$2H_2O$
H x 2				H x 4
O x 2				O x 2
$2H_2$	+	O_2	$\rightarrow$	$2H_2O$
H x 4				H x 4
O x 2				O x 2

Same number of hydrogens and oxygens on both sides – balanced

Moles

The mole is a convenient way of counting atoms, *i.e.* 1 mole of atoms contains 6.02×10^{23} atoms and is used as a unit of measurement.
Gram Formula Mass (GFM) = The mass of 1 mole of atoms/molecules
Gram Formula Mass = Sum of the Relative Atomic Masses for all atoms shown in the chemical formula (page 7 of data booklet)

e.g.

GFM for $CaCO_3$

$$
\begin{array}{l}
CaCO_3 \\
\rightarrow 1 \times 40 = 40 \\
\qquad\qquad\qquad\quad + \\
\rightarrow 1 \times 12 = 12 \\
\qquad\qquad\qquad\quad + \\
\rightarrow 3 \times 16 = 48 \\
\hline
\qquad\quad = 100\text{ g}
\end{array}
$$

Calculations:

Moles equations (page 3 of data booklet)

$$n = \frac{m}{GFM} \qquad \text{moles (mol)} = \frac{\text{moles (mol)}}{\text{GFM (gmol}^{-1}\text{)}}$$

$$n = CV \qquad \text{moles (mol)} = \text{Concentration (mol l}^{-1}\text{)} \times \text{Volume (l)}$$

Volume must be converted to litres, if in cm^3 need to divide by 1000

You must also be able to rearrange these equations to find m, c or v

e.g.

$$m = n \times GFM \qquad C = \frac{n}{V} \qquad V = \frac{n}{C}$$

Calculations from balanced equations

3 step process:

1. Calculate number of moles of the substance been given enough information about
2. Use the mole ratio to calculate the number of moles of the unknown
3. Use the equation again to calculate the mass of the unknown

e.g. 1

The "exploding bottle" experiment reacts hydrogen and oxygen together to produce water.

Calculate the mass of oxygen required to react with excess hydrogen to produce 10 g of water

$$2H_2 + O_2 \rightarrow 2H_2O$$

1. Calculate number of moles:

		$\mathbf{2H_2}$	+	$\mathbf{O_2}$	$\rightarrow$	$\mathbf{2H_2O}$
m				?		10 g
n	=					$\frac{m}{GFM}$
	=					$\frac{10}{(2x1)+(1x16)}$
	=					$\frac{10}{18}$
	=					0.556 moles

2. Use the mole ratio to calculate the number of moles of the unknown (÷ 2 here as 2:1 mole ratio).

		$\mathbf{2H_2}$	+	$\mathbf{O_2}$	$\rightarrow$	$\mathbf{2H_2O}$
n	=			0.278 moles	$\xleftarrow{\div 2}$	0.556 moles

3. Use the equation again to calculate the mass of the unknown

		$\mathbf{2H_2}$	+	$\mathbf{O_2}$	$\rightarrow$	$\mathbf{2H_2O}$
m	=			nxGFM		
				0.278x(16x2)		
				0.78x2		
				8.89 g		

e.g. 2

In a titration 20 cm^3 of KOH was needed to neutralise 15 cm^3 of H_2SO_4 (concentration 0.1 mol l^{-1}). Calculate the concentration of KOH.

$$2KOH + H_2SO_4 \rightarrow K_2SO_4 + 2H_2O$$

1. Calculate number of moles:

		2KOH	**+**	**H_2SO_4**	**→**	**K_2SO_4**	**+**	**$2H_2O$**
C		?		0.1 mol l^{-1}				10 g
V		20 cm^3		15 cm^3				
	÷ 1000 for litres	0.02 l		0.015 l				

n = C x V

= 0.1 x 0.015

= 0.0015 moles

2. Use the mole ratio to calculate the number of moles of the unknown (x 2 here as 1:2 mole ratio).

$$2KOH + H_2SO_4 \rightarrow K_2SO_4 + 2H_2O$$

n = 0.0030 moles $\xleftarrow{\times 2}$ 0.0015 moles

3. Use the equation again to calculate the mass of the unknown

$$2KOH + H_2SO_4 \rightarrow K_2SO_4 + 2H_2O$$

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

$$= \frac{0.003}{0.02}$$

<u>0.15 mol l^{-1}</u>

Calculations involving: $n = CV$ and $n = \frac{m}{GFM}$

e.g.

5 g of sodium hydroxide (NaOH) is dissolved in 250 cm^3 of water. Calculate the concentration of the sodium hydroxide.

moles of NaOH = $\frac{m}{GFM}$ $\qquad$ **$C = \frac{n}{V}$**

moles of NaOH = $\frac{5}{40}$ $\qquad$ $C = \frac{0.125}{0.25}$

moles of NaOH = 0.125 mol $\qquad$ **<u>0.5 mol l^{-1}</u>**

Percentage composition

$$\% \text{ by mass} = \frac{m}{GFM} \times 100 \qquad \% \text{ by mass} = \frac{\text{mass of element in formula}}{GFM} \times 100$$

e.g.

Calculate the percentage of nitrogen in ammonium nitrate (NH_4NO_3)

NH_4NO_3

→ 2 x 14 = 28 +

→ 4 x 1 = 4 +

→ 3 x 16 = 48

= 80 g

$$\% \text{ by mass} = \frac{28}{80} \times 100 = \underline{\mathbf{35.0\ \%}}$$

(d) Acids and bases

(i) pH

- acid below 7 (universal indicator: red, orange, yellow)
- neutral 7 (universal indicator: green)
- base above 7 (universal indicator: blue, purple)

Indicators = universal indicator, pH paper, phenolphthalein, litmus, etc.
pH of a solution is related to the concentration of hydrogen ions (H^+) in solution

- Diluting an **acid** makes its **pH increase** towards 7 (less acidic)
    - the concentration of hydrogen ions (H^+) is decreased.
- Diluting an **alkali** makes its **pH decrease** towards 7 (less alkaline)
    - the concentration of hydroxide ions (OH^-) is decreased.

To change the pH by 1, a 10 fold dilution is required (*e.g.* adding 9 cm^3 of water to 1 cm^3 of acid/alkali)

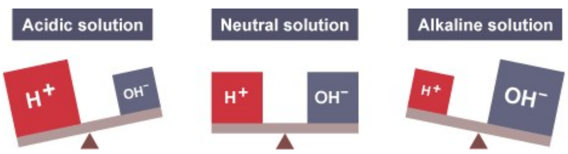
Dissociation:

Water naturally breaks down (dissociates) to form positive hydrogen ions (H^+) and negative hydroxide ions (OH^-).

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

This is a reversible reaction and a very small proportion of water molecules dissociate.
The small number of charged ions means water is able to **conduct electricity**.

- Neutral solutions (e.g. water) contain an **<u>equal</u>** concentration (number of) of hydrogen (H^+) and hydroxide ions (OH^-)
- All acidic solutions are a source of hydrogen ions (H^+), they contain a **greater concentration of hydrogen ions (H^+)** than hydroxide ions (OH^-)
- All alkaline solutions contain **more hydroxide ions (OH^-)** than hydrogen ions (H^+)



Common laboratory acids and alkalis [An alkali is a soluble base]

Acid Name	Formula	Dissociates in water as…
Hydrochloric Acid	HCl	H^+ (aq) + Cl^- (aq)
Sulphuric Acid	H_2SO_4	$2H^+$ (aq) + SO_4^{2-} (aq)
Nitric Acid	HNO_3	H^+ (aq) + NO_3^- (aq)

Alkali Name	Formula	Dissociates in water as…
Sodium hydroxide	$NaOH$	Na^+ (aq) + OH^- (aq)
Calcium hydroxide	$Ca(OH)_2$	Ca^{2+} (aq) + $2OH^-$ (aq)

Conductivity: Both acidic and alkaline solutions can conduct electricity because they contain ions that are charged and free to move in solution

Formation of acids and alkalis from oxides:

- When added to water, **soluble** **metal oxides** form an **alkaline** solution (i.e. they produce metal hydroxide solutions, e.g. $CaO + H_2O \rightarrow Ca(OH)_2$, causing the pH to increase)
- **Soluble** **non-metal oxides** (e.g. SO_2 and NO_2) will dissolve in water to form **acidic** solutions (i.e. increase the hydrogen ion concentration) [e.g. acid rain]
- If oxides are **insoluble** they will NOT change the pH of water and it will remain neutral e.g. silver oxide is insoluble

Use page 8 of the data booklet to determine if something is soluble or insoluble

Neutralisation reactions

The equation for neutralisation is:

Acid + Base → Salt + Water

A base is any substance that neutralises an acid, an alkali is a **soluble** base therefore:

Acid + Alkali → Salt + Water

e.g.

Hydrochloric acid + Potassium Hydroxide → Potassium Chloride + Water

Metal oxides, metal hydroxides, metal carbonates and ammonia are all examples of **bases**:

Acid + Metal oxide → Salt + Water

Acid + Metal hydroxide → Salt + Water

Acid + Metal carbonate → Salt + Water + Carbon dioxide

Acid + ammonia → ammonium salt + water

Salts: Metals can react with acids to form a salt however instead of water, hydrogen is formed:

Acid + Metal → Salt + Hydrogen

Naming of the salt:

The salt formed is the acid with the hydrogen ion replaced by a metal ion, e.g. **sodium** hydroxide and hydro**chloric** acid makes water and **sodium chloride**.
i.e. the first part of the name comes from the **base** and the second part comes from the **acid**

e.g. **copper** oxide + **sulphuric** acid → **copper** **sulphate** + water

Acid Name	**Salt name ends in**
Phosphoric acid	Phosphate
Hydrochloric acid	Chloride
Sulphuric acid	Sulphate
Nitric acid	Nitrate

Neutralisation reactions can be used to prepare soluble salts

Separation of salts from water:

Soluble salts – separate by evaporation

Insoluble salts – separate by filtration

Insoluble salts can be made from other soluble salts by a **precipitation reaction.** Two solutions react to form the **insoluble** product called the **precipitate** which can be **filtered** off

e.g.

Magnesium **sulphate** + **Barium** chloride → **Barium** **sulphate** + Magnesium chloride

Magnesium chloride is soluble and barium sulphate is insoluble therefore barium sulphate is the precipitate *(Use page 8 of your data booklet)*

Equations for Neutralisations:

Word Equation

Hydrochloric Acid + Sodium Hydroxide → Sodium Chloride + Water

Chemical Equation

$HCl\,(aq) + NaOH\,(aq) \rightarrow NaCl\,(aq) + H_2O\,(l)$

Ionic Equation:

$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-\,(aq) + H_2O(l)$

In this type of reaction there are some ions that do not participate in the reaction, we call these **spectator ions**. They appear on both sides of the reaction arrow and are unchanged.

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$

Ionic equation without spectator ions (removed Cl^- and Na^+):

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Titration

Titration is an analytical technique used to determine the accurate volumes involved in chemical reactions (*e.g.* neutralisations)
An indicator is used to show the end point of the reaction.
A **pipette** is used to accurately measure out a volume of alkali and a **burette** is used to add the volume of acid accurately
White tile is used to help accurately show the colour change and endpoint
Rough titration gives an approximate end point – do NOT use for average
Average calculated from two concordant results, *i.e.* volumes within 0.2 cm^3
e.g. 10.2 cm^3 and 10.4 cm^3 would be concordant giving an average of 10.3 cm^3
p