



PURE

CHEMISTRY

SEC 4

NOTES

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The Periodic Table of Elements

Group																	
I	II	1 H hydrogen 1										III	IV	V	VI	VII	0
		<div>Key</div> <div>proton (atomic) number atomic symbol name relative atomic mass</div>															
3 Li lithium 7	4 Be beryllium 9																
11 Na sodium 23	12 Mg magnesium 24																
19 K potassium 39	20 Ca calcium 40	21 Sc scandium 45	22 Ti titanium 48	23 V vanadium 51	24 Cr chromium 52	25 Mn manganese 55	26 Fe iron 56	27 Co cobalt 59	28 Ni nickel 59	29 Cu copper 64	30 Zn zinc 65	5 B boron 11	6 C carbon 12	7 N nitrogen 14	8 O oxygen 16	9 F fluorine 19	10 Ne neon 20
37 Rb rubidium 85	38 Sr strontium 88	39 Y yttrium 89	40 Zr zirconium 91	41 Nb niobium 93	42 Mo molybdenum 96	43 Tc technetium -	44 Ru ruthenium 101	45 Rh rhodium 103	46 Pd palladium 106	47 Ag silver 108	48 Cd cadmium 112	49 In indium 115	50 Sn tin 119	51 Sb antimony 122	52 Te tellurium 128	53 I iodine 127	54 Xe xenon 131
55 Cs caesium 133	56 Ba barium 137	57 – 71 lanthanoids	72 Hf hafnium 178	73 Ta tantalum 181	74 W tungsten 184	75 Re rhenium 186	76 Os osmium 190	77 Ir iridium 192	78 Pt platinum 195	79 Au gold 197	80 Hg mercury 201	81 Tl thallium 204	82 Pb lead 207	83 Bi bismuth 209	84 Po polonium -	85 At astatine -	86 Rn radon -
87 Fr francium -	88 Ra radium -	89 – 103 actinoids	104 Rf Rutherfordium -	105 Db dubnium -	106 Sg seaborgium -	107 Bh bohrium -	108 Hs hassium -	109 Mt meitnerium -	110 Ds darmstadtium -	111 Rg roentgenium -	112 Cn copernicium -		114 Fl flerovium -		116 Lv livermorium -		

lanthanoids

57 La lanthanum 139	58 Ce cerium 140	59 Pr praseodymium 141	60 Nd neodymium 144	61 Pm promethium -	62 Sm samarium 150	63 Eu europium 152	64 Gd gadolinium 157	65 Tb terbium 159	66 Dy dysprosium 163	67 Ho holmium 165	68 Er erbium 167	69 Tm thulium 169	70 Yb ytterbium 173	71 Lu lutetium 175
89 Ac actinium -	90 Th thorium 232	91 Pa protactinium 231	92 U uranium 238	93 Np neptunium -	94 Pu plutonium -	95 Am americium -	96 Cm curium -	97 Bk berkelium -	98 Cf californium -	99 Es einsteinium -	100 Fm fermium -	101 Md mendelevium -	102 No nobelium -	103 Lr lawrencium -

actinoids

The volume of one mole of any gas is 24 dm^3 at room temperature and pressure (r.t.p.).



4EXP CHEMISTRY TOPICS

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The Periodic Table is an arrangement and classification of the elements.

It can be used to predict the properties of elements from trends that are observed.

 metals  non-metals  metalloids

- Vertical columns of elements in the Periodic Table
- The Group number indicates the number of valence electrons in an atom of an element.
- Element in the same group have the same number of valence electrons, hence they form ions of the same charges and have similar chemical properties (reaction)

Element	Electronic Configuration	Valence electrons	Group	Ions
Lithium	2.1	1	I	Li ⁺
Sodium	2.8.1	1	I	Na ⁺
Potassium	2.8.8.1	1	I	K ⁺
Oxygen	2.6	6	VI	O ²⁻
Sulphur	2.8.6	6	VI	S ²⁻
Beryllium	2	2	II	Be ²⁺
Magnesium	2.8.2	2	II	Mg ²⁺



Exception

Element	Electronic Configuration	Valence electrons	Group
Helium	2	2	0

Helium is not located in Group II, as its 1st shell is completely full and it has a stable noble gas configuration.

Trends down a Group

The following trends are observed when down a Group:

- The proton number becomes bigger.
The number of energy levels increases as you move down a group as the number of electrons increases. Each subsequent energy level is further from the nucleus than the last. Therefore, the atomic radius increases as the group and energy levels increase.
- The properties of the elements become more metallic. This is because on going down the group, the elements lose electrons more easily, as the valence electrons become spaced further away from the positive nucleus.

PERIOD (Horizontal row)

- Horizontal rows of elements in the Periodic Table.
- The Periodic Table is a list of elements arranged in order of their increasing atomic or proton number
- The Period number indicates the number of occupied electron shells in an atom of an element.



Trends across a Period

- Across the period, elements changes from **Metals to Non-metals**
 - A decrease in metallic properties
 - An increase in non-metallic properties
 - This results in the different types of oxides formed.

Group	I	II	III	IV	V	VI	VII	0
Symbol	Na	Mg	Al	Si	P	S	Cl	Ar
Properties	Metallic			Metalloid (Elements having the properties of both metals and non- metals)	Non-metallic			
Nature of oxide	Basic oxide		Amphoteric oxide	Acidic oxide				

Decreasing metallic properties

Increasing non-metallic properties

- The atomic radius decreases. As you go across a period, electrons are added to the same energy level. At the same time, protons are being added to the nucleus. The concentration of more protons in the nucleus creates a "**higher effective nuclear charge**." In other words, there is a stronger force of 'positive attraction' pulling the electrons closer to the nucleus resulting in a smaller atomic radius.
- The electronegativity increases (electronegativity is a measure of the ability of an atom or molecule to attract pairs of electrons in a chemical bond.)
- The tendency for an element to gain electrons to form negative ions, increases across the period.



Difference in Chemical properties between

Metals	Non Metals
Forms basic oxides	Forms acidic oxides (or neutral oxide like H_2O , CO)
Tend to lose electrons to gain a positively charged ion	Tend to gain electrons to form a negatively charged ion

Difference in Physical properties between

Metals	Non Metals
Good conductor of heat and electricity	Poor conductor of heat and electricity
Generally malleable, ductile	Generally brittle if solid
Solid at room temperature (except Mercury)	Generally liquid or gas at room temperature, they can exist as solids too.

- The metallic elements between Gp II and III are known as the **Transition Metals**

Transition metals have:

- High electrical conductivity due to delocalized/free flowing electrons
- Varied oxidation states eg. Fe^{2+} , Fe^{3+} , Cu^+ , Cu^{2+}
- Formed compounds that are coloured.

Compound	Colour	Compound	Colour
Iron (II) hydroxide	Green	Copper (II) oxide	Black
Iron (III) hydroxide	Brown	Copper (II) carbonate	Green
Iron (III) chloride	Yellow/Orange	Copper (II) sulfate	Blue
Potassium Manganate (VII)	Purple	Manganese (IV) oxide	Black



GROUP I: ALKALI METALS

Physical properties of Group I

- Shiny and silvery in appearance when freshly cut
- Are soft can be cut easily with a knife
- Good conductors of heat and electricity
- Have low melting and boiling point (compared to other metals) – Mp & Bp decreases down Gp I
- Have low densities (eg. Li, Na and K can float on water) – Density Increases down Gp I

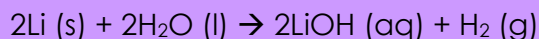


Li Lithium
Na Sodium
K Potassium
Rb Rubidium
Cs Caesium
Fr Francium

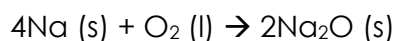
Chemical properties of Group I

- Have one electron in its valence shell. Hence similar chemical properties
- Highly reactive metals - Reactivity increases down the group
- Soluble in water to form alkalis

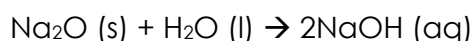
Group I metals + H₂O → Metal hydroxide + Hydrogen gas



- React readily with oxygen in air to form basic oxides



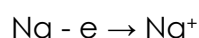
The basic oxides formed from Group I metals can further dissolve in water to give alkaline solution. Alkaline solution will turn red litmus blue and universal indicator purple



Hence, these reactive metals are stored in oil to prevent itself from reacting with air and moisture.



- Alkali metals are powerful reducing agents as it loses its valence electrons easily to attain a stable noble gas structure.



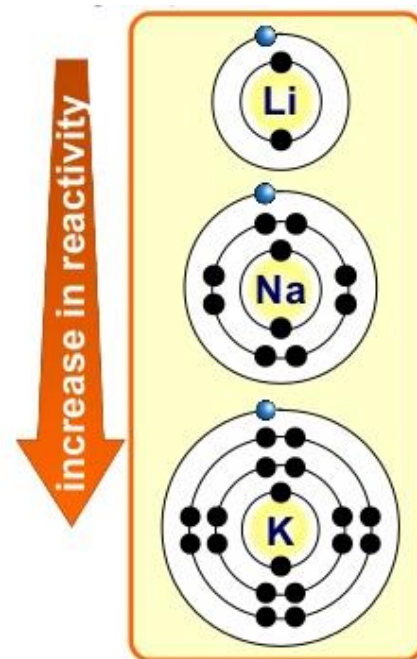
Trend in Chemical Properties of Group I

Down Group I, the reactivity and reducing power of the metals increases.

Reason: As size of atom increases, ease of losing electrons from bigger atoms increases. In the larger atoms, the valence electron is further away from the nucleus. The nuclear charge holding the valence electron to the atom is weak. Hence it is more easily lost.

Compounds of Group I

- Ionic compounds - Lattice structure
- Soluble in water - Alkali (Universal indicator will turn purple)
- Similar chemical formula



Element	Chloride	Hydroxide (Alkali)	Sulfate
Li	LiCl	LiOH	Li ₂ SO ₄
K	KCl	KOH	K ₂ SO ₄
Rb	RbCl	RbOH	Rb ₂ SO ₄



GROUP VII: HALOGENS

Physical properties of Group VII

- Have low melting point and boiling points - bp and mp increases down the group.
- Are coloured - The colour of the halogen becomes darker down the group.

Element	Colour	State at room temperature
F ₂	Pale yellow	Gas
Cl ₂	Greenish yellow	Gas
Br ₂	Reddish brown	Liquid
I ₂	Purplish black	Solid
At ₂	Black	Solid

F

Fluorine

Cl

Chlorine

Br

Bromine

I

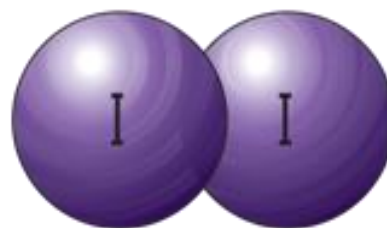
Iodine

At

Astatine

Chemical properties of Group VII

- Are reactive non-metals - They are also said to be electronegative as they have a high tendency to gain an additional electron in its valence shell to achieve a more stable octet configuration.



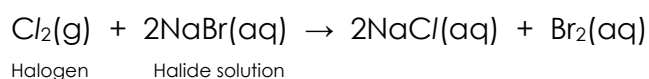
They do so by sharing an electron (in the formation of covalent compounds) or by accepting an electron (in the formation of an ionic compound).

- Halogens reacts with most metals to form salts called halides. F⁻, Cl⁻, Br⁻ etc are examples of halides ions. These are ionic compounds.



- Have seven electrons in their valence shell
- They exists as diatomic, covalent molecules
- Undergo **displacement reaction** with more reactive halogen.

The more reactive halogen can displace the less reactive halogen from its halide solution. Example: Chlorine gas is able to displace bromine from sodium bromide to form bromine and sodium chloride.

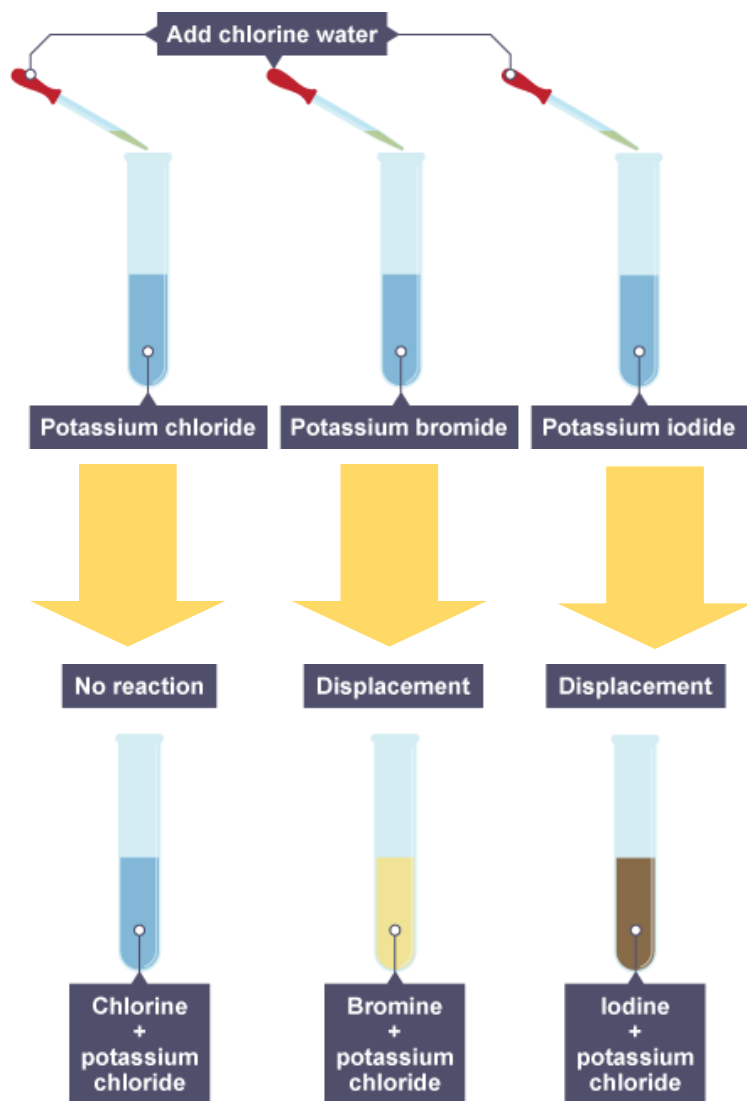




- Bromine is less reactive than chlorine. It cannot displace chlorine from potassium chloride solution. A less reactive halogen cannot displace a more reactive halogen from its halide solution.

More examples of displacement reaction

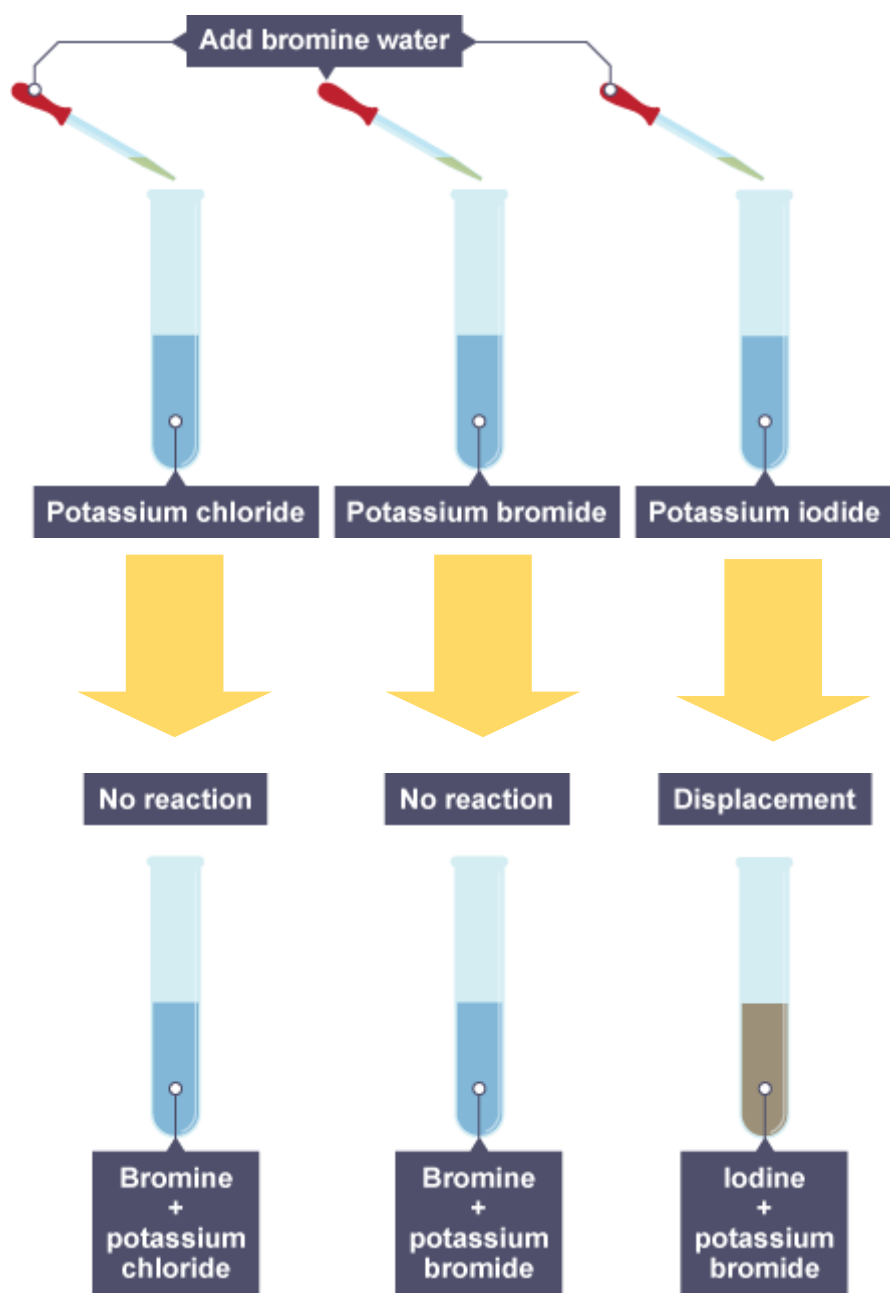
Example 1: Chlorine water is added to three solutions



Chemical Equations	Observation
$\text{Cl}_2 (\text{aq}) + \text{KCl} (\text{aq}) \rightarrow \text{no reaction}$	KCl solution remains colourless.
$\text{Cl}_2 (\text{aq}) + 2\text{KBr} (\text{aq}) \rightarrow 2\text{KCl} (\text{aq}) + \text{Br}_2 (\text{aq})$	Colourless KBr solution turns brown due to Br_2
$\text{Cl}_2 (\text{aq}) + 2\text{KI} (\text{aq}) \rightarrow 2\text{KCl} (\text{aq}) + \text{I}_2 (\text{aq})$	Colourless KI solution turns brown due to I_2 . Iodine when dissolved is brown.



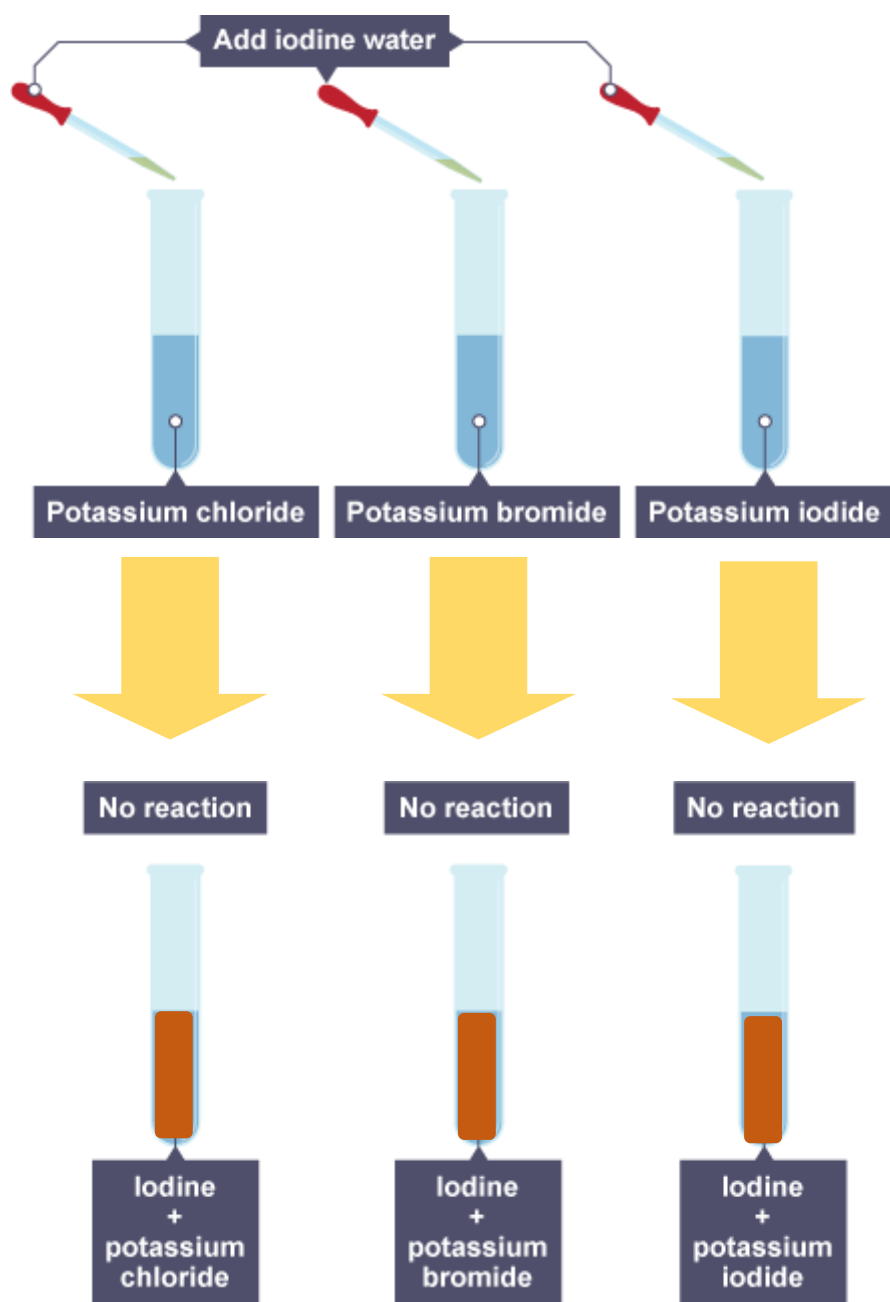
Example 2: Aqueous Bromine is added to three solutions



Chemical Equations	Observation
$\text{Br}_2 (\text{aq}) + \text{KCl} (\text{aq}) \rightarrow \text{no reaction}$	KCl solution remains colourless.
$\text{Br}_2 (\text{aq}) + \text{KBr} (\text{aq}) \rightarrow \text{no reaction}$	KBr solution remains colourless.
$\text{Br}_2 (\text{aq}) + 2\text{KI} (\text{aq}) \rightarrow 2\text{KBr} (\text{aq}) + \text{I}_2 (\text{aq})$	Colourless KI solution turns brown due to I_2 . Iodine when dissolved is brown.



Example 2: Aqueous iodine is added to three solutions

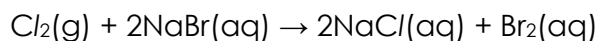


Chemical Equations	Observation
$I_2 (aq) + KCl (aq) \rightarrow \text{no reaction}$	KCl, KBr and KI remains colourless. There is no displacement. The mixture is brown due to the unreacted I_2 .
$I_2 (aq) + KBr (aq) \rightarrow \text{no reaction}$	
$I_2 (aq) + KI (aq) \rightarrow \text{no reaction}$	

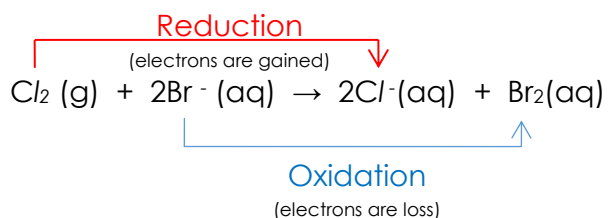


Additional Notes (Redox Reaction)

Displacement reaction are redox reaction



Ionic equation:



Redox is explained in terms of electron gain and loss.

The equation above can be separated into two parts:



Trends in Chemical Properties in Group VII

Down the group,

- Reactivity decreases;
- Oxidising power decreases.
- As size of atom increases, it becomes more difficult for the nucleus to draw an additional electron into its valence shell.



Iodine at room temperature



GROUP 0 - NOBLE GASES

Physical Properties

Group 0 elements are known as noble gases

- Colourless and odourless gases at room temperature
- Very low melting points and boiling points
- Are insoluble in water

Chemical Properties

Element	Electronic Configuration	
Helium	2	(Duplet configuration)
Neon	2.8	(Octet configuration)
Argon	2.8.8	(Octet configuration)

Noble gases have stable electronic structures. Hence, they

- are chemically unreactive;
- do not gain, lose or share electrons;
- do not form compounds;
- exist as monatomic gases.

Uses of Noble gases

Element	Uses
He	Use to fill weather or advertisement balloons and airships Use by divers in breathing tanks in preference to Nitrogen
Ne	Use in making colourful light boards and signages.
Ar	Use to fill up light bulb to prevent filament from oxidation Use for welding of stainless steel

He Helium
Ne Neon
Ar Argon
Kr Krypton
Xe Xenon
Rn Radon





TRANSITION METALS

Properties of transition metals

Metals located in the block between Gp II and Gp III

Have high melting and boiling point

Have high densities

These metals have more than one oxidation state (charge) for the different compounds formed eg

Metal	Compound	Oxidation state of metal
Copper	CuCO_3	Charge of Cu is +2 (Cu^{2+})
	Cu_2O	Charge of Cu is +1 (Cu^+)
	CuO	Charge of Cu is +2 (Cu^{2+})
Iron	FeSO_4	Charge of Fe is +2 (Fe^{2+})
	Fe(OH)_3	Charge of Fe is +3 (Fe^{3+})
Manganese	KMnO_4	Charge of Mn is +7 (Mn^{7+})
	MnO_2	Charge of Mn is +4 (Mn^{4+})

Formed coloured compounds -





Colours of the compounds of a transition metal are different at different oxidation states.





Colour of hydrated compounds may differ from anhydrous compound.

Compounds of transition metals are used in dyes, pigments and paints to produce different colours

Compound	Colour of compound	Compound	Colour of compound
CuCO_3	Green	FeSO_4	Green
CuO	Black	Fe(OH)_3	Brown
CuSO_4	Blue	KMnO_4	Purple



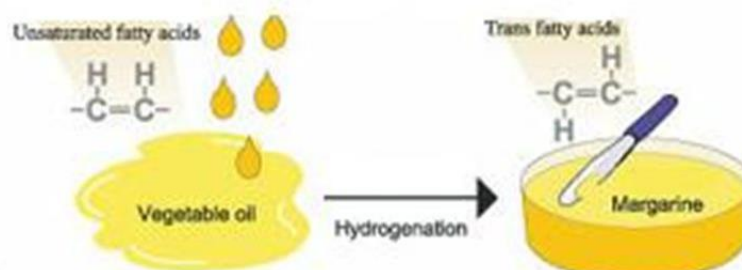
	Chromium (Cr)		Iron (Fe)	
Compound				
	Chromium(III) chloride	Potassium dichromate(VI)	Iron(II) sulfate	Iron(III) chloride
Chemical formula	CrCl_3	$\text{K}_2\text{Cr}_2\text{O}_7$	FeSO_4	FeCl_3
Oxidation state	+3	+6	+2	+3
Colour	Green	Orange	Pale green	Yellow

	Chromium (Cr)		Iron (Fe)	
Compound				
	Chromium(III) chloride	Potassium dichromate(VI)	Iron(II) sulfate	Iron(III) chloride
Chemical formula	CrCl_3	$\text{K}_2\text{Cr}_2\text{O}_7$	FeSO_4	FeCl_3
Oxidation state	+3	+6	+2	+3
Colour	Green	Orange	Pale green	Yellow

Transition metals and their compounds are good catalysts.

A catalyst is a substance that increases the speed of a chemical reaction and remains chemically unchanged at the end of the reaction.

Catalyst	Reaction / Process
Manganese, Mn	Decomposition of hydrogen peroxide to water and oxygen gas
Iron, Fe	Manufacture of ammonia gas in the Haber process
Nicke, Ni	Hydrogenation of vegetable oil to margarine





Rate of Reaction

What is the Rate of reaction?

- The **speed** of a reaction is called the **rate** of the reaction.
- Rate of reaction is the rate at which a **chemical reaction takes place**.

PART I: MEASURING SPEED OF REACTION

The speed of the reaction or rate can be measured by:

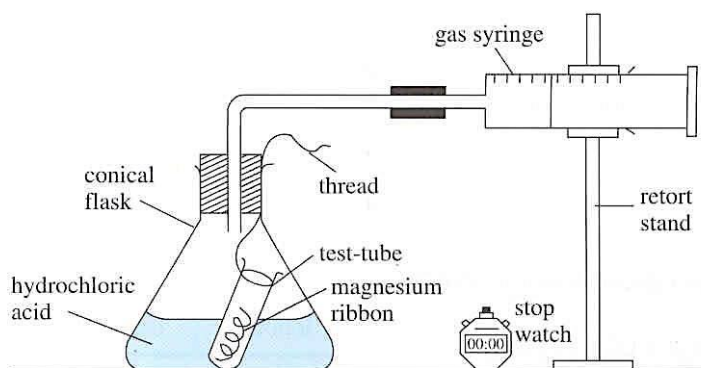
- Volume of gas produced per unit time
- Mass of reactant used up per unit time
- Rate of effervescence / bubbling (Not accurate)

(a) Measuring speed of reaction through volume of gas produced.

The speed of a reaction can be found by measuring the volume of gas produced measured at regular time intervals.

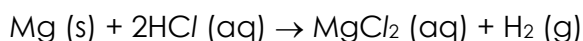
EXPERIMENT 1:

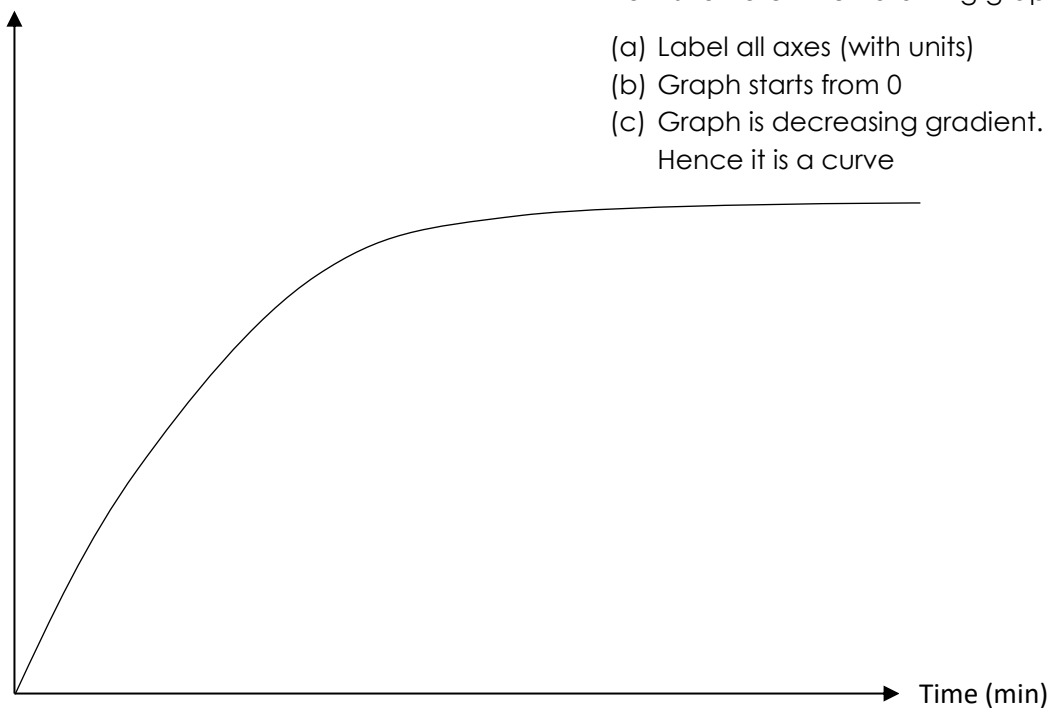
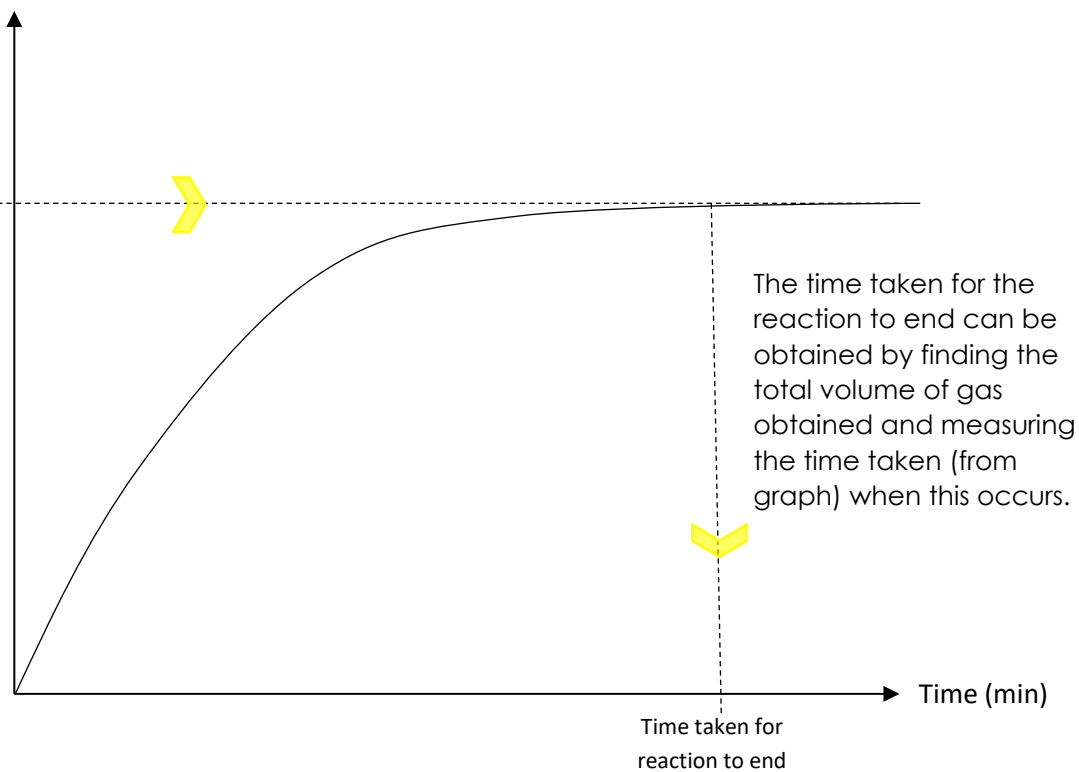
The setup is shown below. A **fixed mass** of magnesium metal was added to an excess of dilute hydrochloric acid. The magnesium metal was entirely dissolved and volume of hydrogen gas evolved was measured at regular time interval during the reaction.

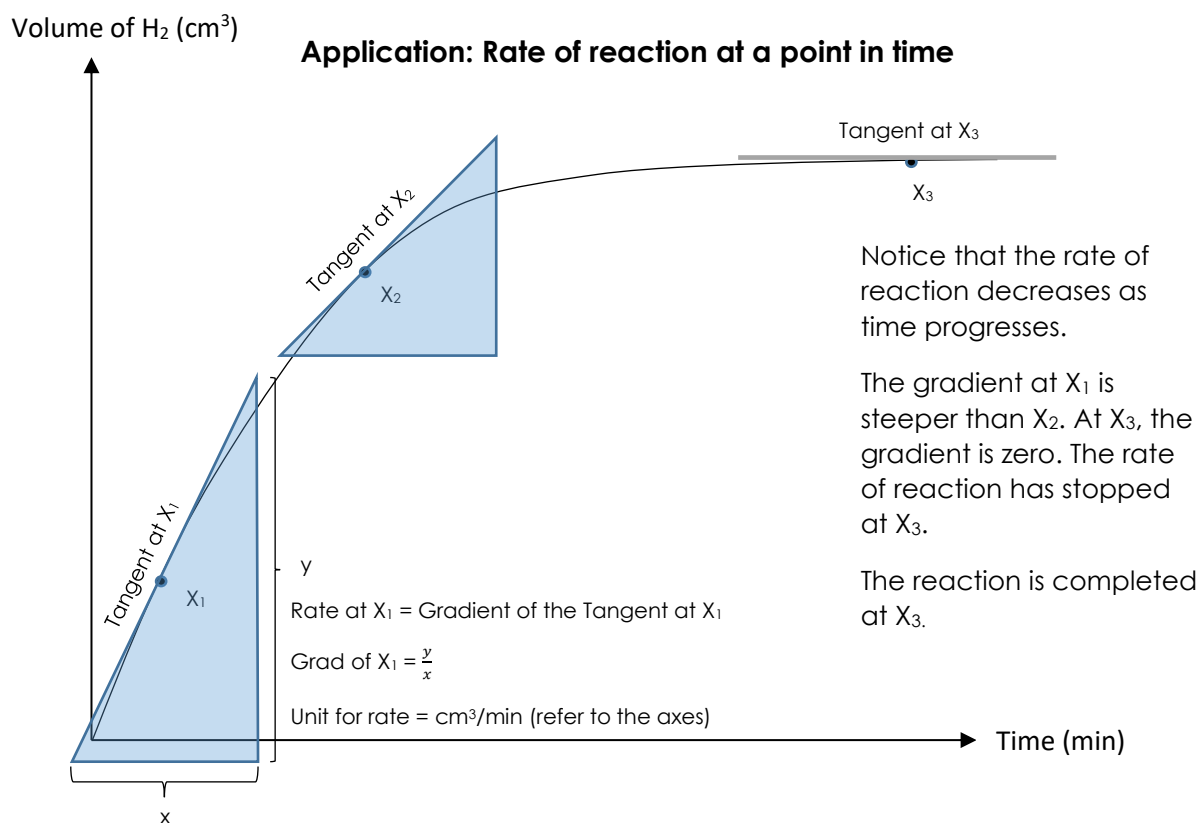
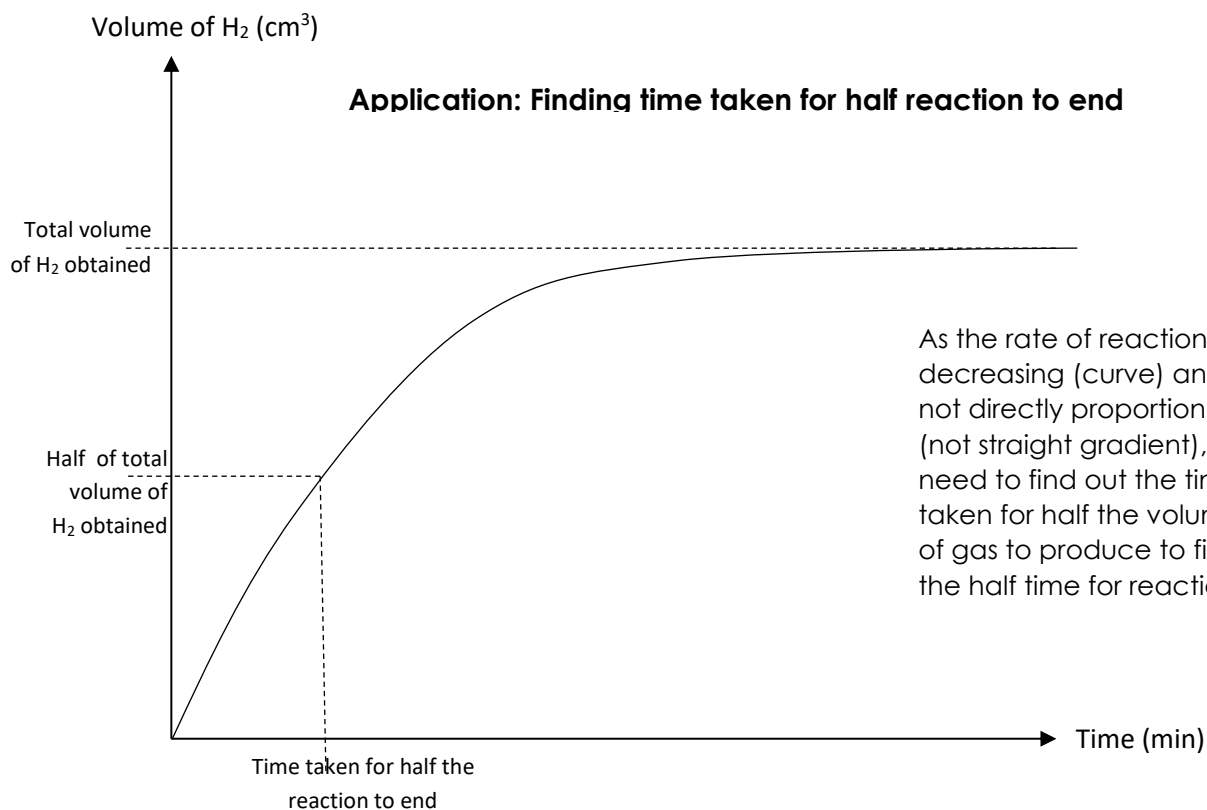


Time (min)	Volume of Hydrogen Measured (cm ³)
0	0
2	10
4	20
6	25
8	30
10	30
12	30

The chemical equation of the reaction is shown below.



**RESULTS**Volume of H_2 (cm^3)**GRAPH INTERPRETATION & APPLICATIONS**Volume of H_2 (cm^3)**Application: Finding time taken for reaction to end**Total volume
of H_2 obtained





(b) Measuring speed of reaction through the loss of mass

The speed of a reaction can be found by measuring these quantities at regular time intervals:

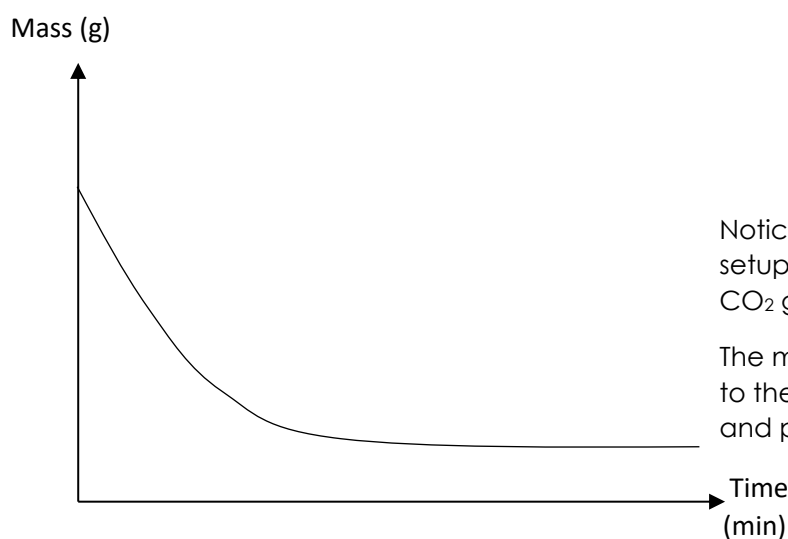
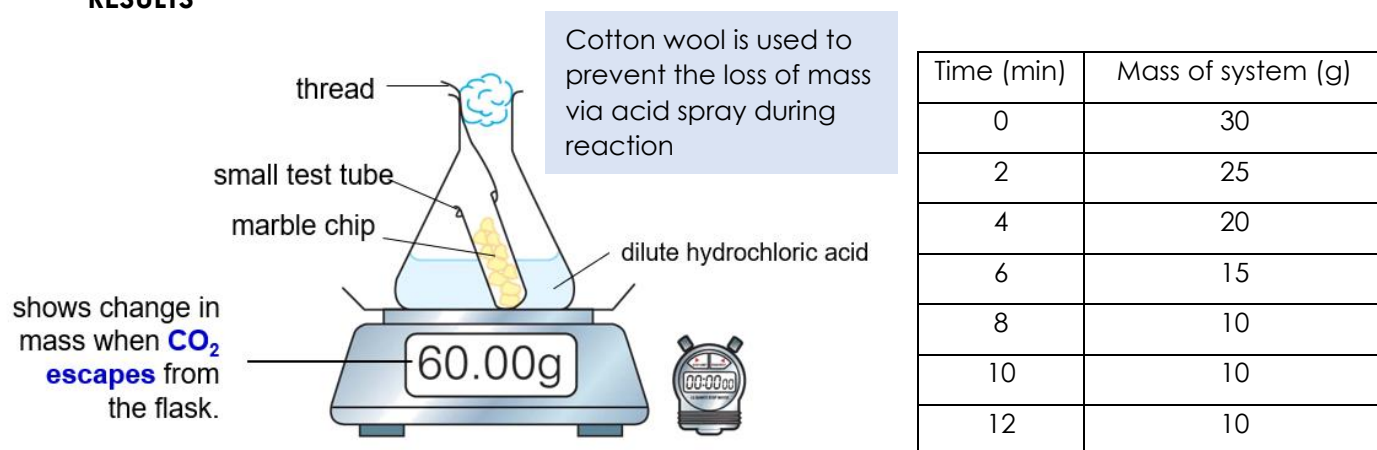
- The change in mass of the system
- As gas is formed during the reaction, it gradually escapes and the overall mass of the system decreases.

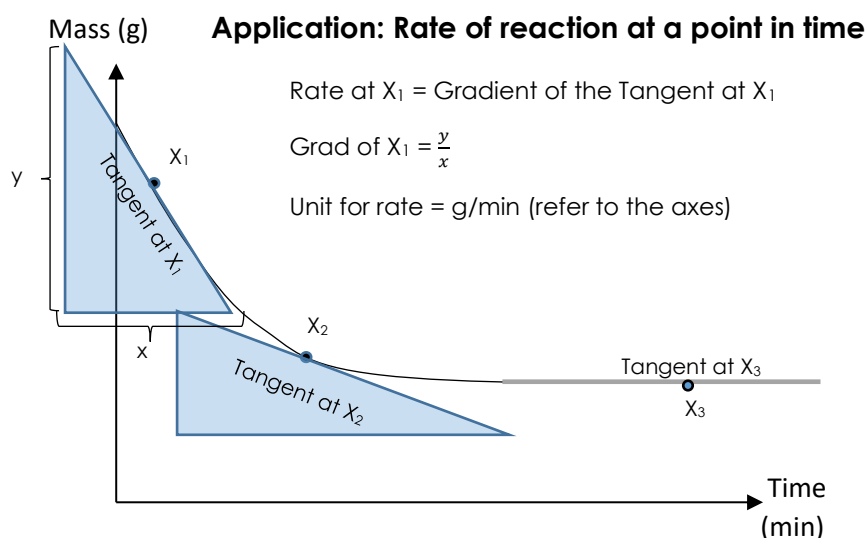
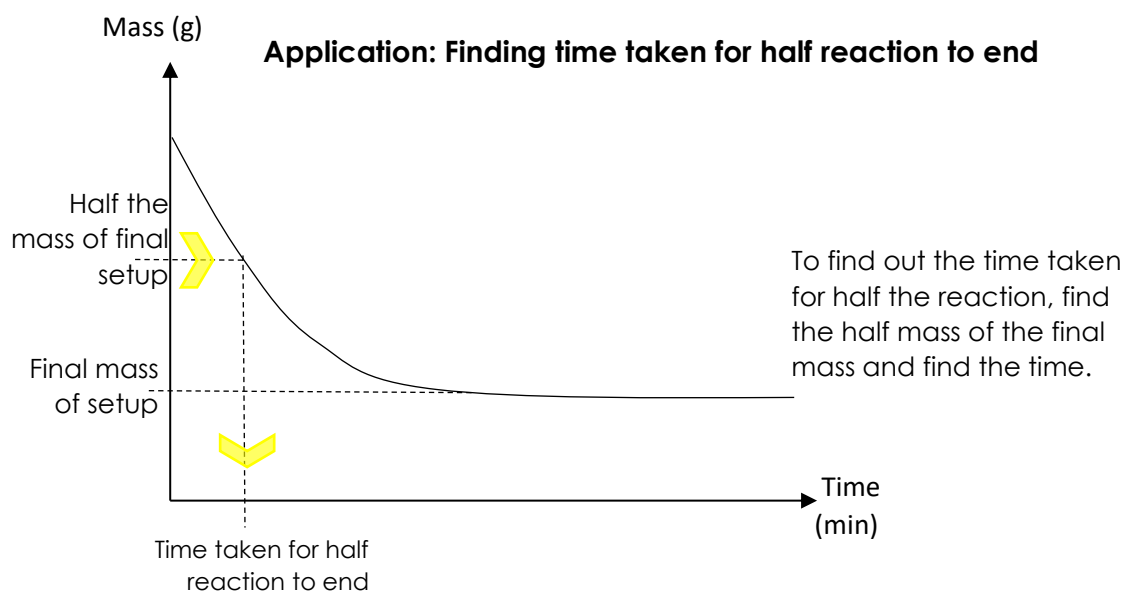
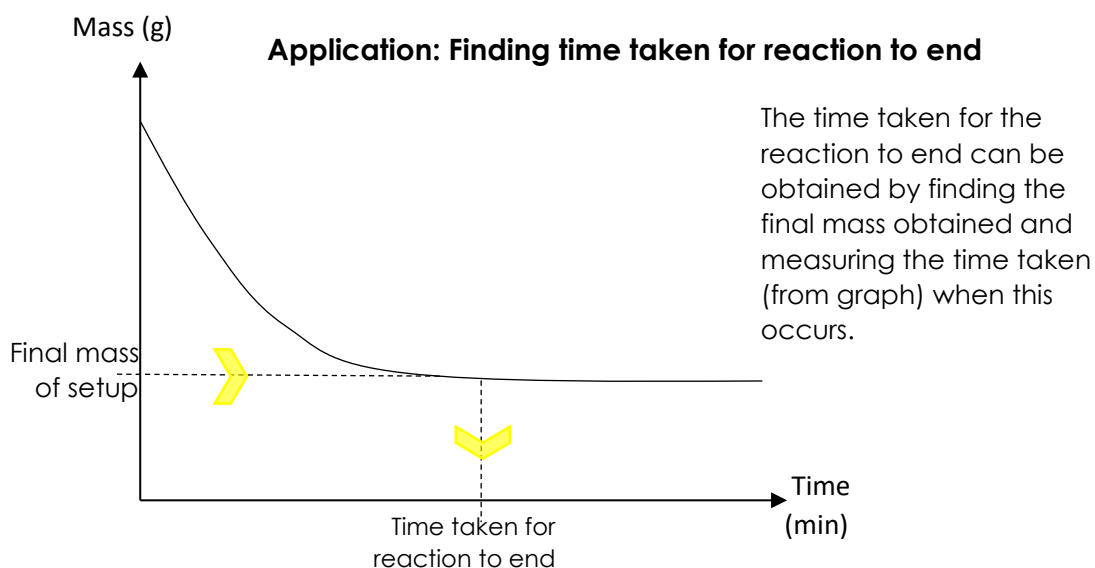
EXPERIMENT 1:

The setup is shown below. A fixed mass of calcium carbonate (marble chip) was added to excess dilute hydrochloric acid in a conical flask. The solid calcium carbonate was entirely dissolved and the volume of carbon dioxide gas evolved was measured at regular time interval during the reaction



RESULTS





Notice that the rate of reaction decreases as time progresses.

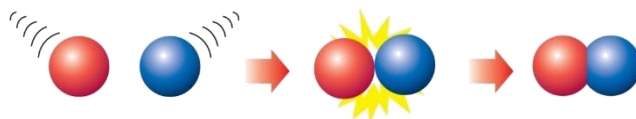
The gradient at X_1 is steeper than X_2 . At X_3 , the gradient is zero. The rate of reaction has stopped at X_3 .

The reaction is completed at X_3 .



PART II: COLLISION THEORY TO EXPLAIN RATE OF REACTION

The rate of reaction can be explained based on the movement, energy and frequency of effective collisions of reactant particles to form products.



This form of explanation based on particles is called Collision Theory

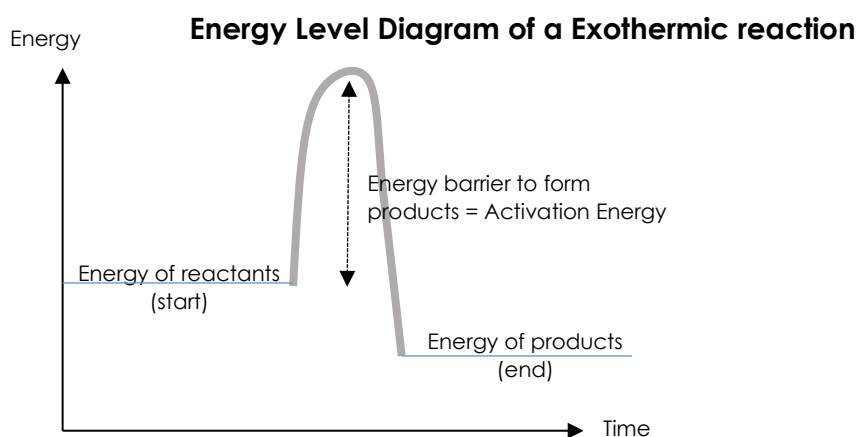
FACTORS AFFECTING THE RATE OF REACTION

The speed of a reaction depends on the following factors:

- (a) the **Concentration** of the reactants
- (b) the **Pressure** of the reactants (for gaseous reactants only)
- (c) the **Particle size** or total **Surface area** of the reactants
- (d) the **Temperature** at which the reaction is occurring.
- (e) **Catalyst** present

For a reaction to occur between two particles:

- the reacting particle must collide with each other
- they must collide with a certain minimum amount of energy known as activation energy

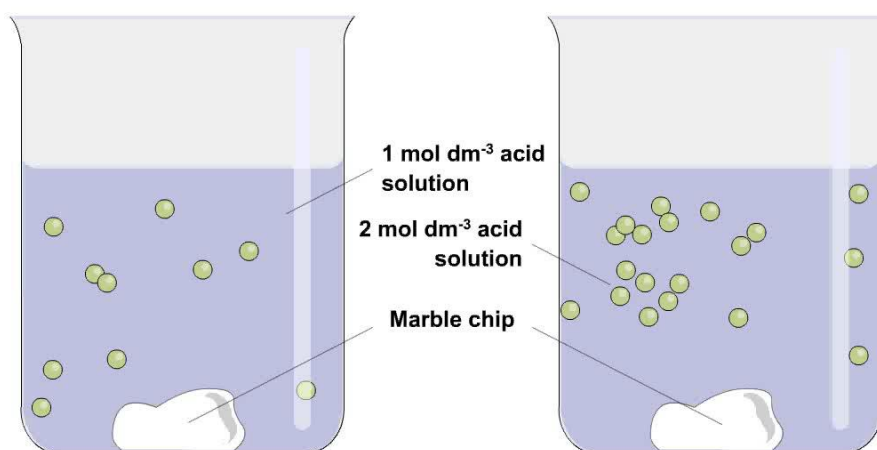


The reactant particles must have enough energy (equal to or higher than the activation energy) in order to collide **effectively** with each other to form the products.

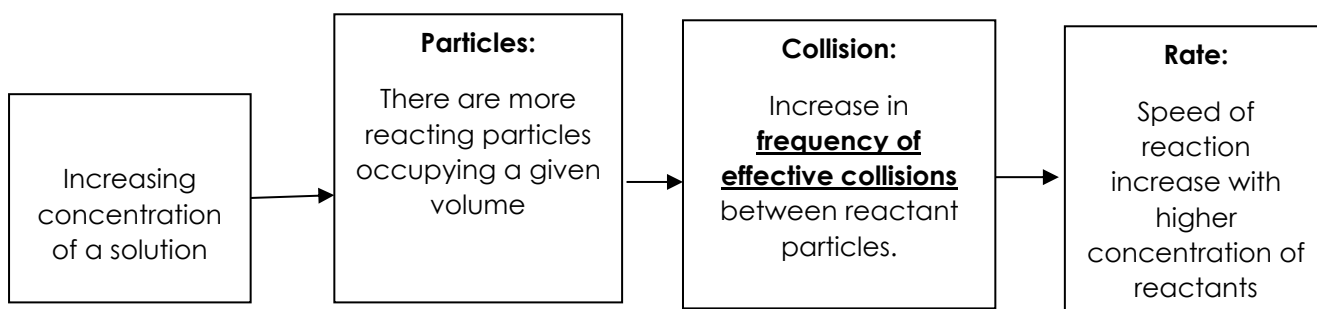
- To form a product, the collisions of reactants must be effective.
- Ineffective collisions will not form products. This happens when the particles do not have enough energy (activation energy)



(a) Concentration of the reactants



FLOW CHART EXPLANATION

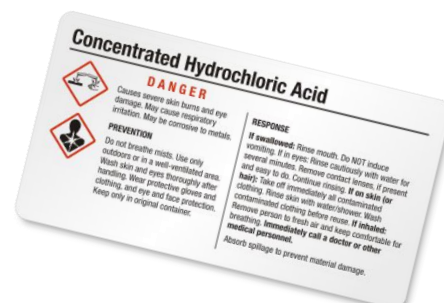
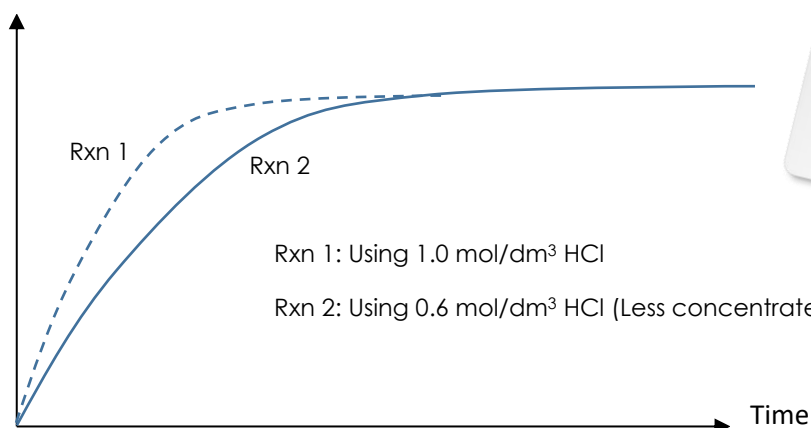


EXPLANATION

Increasing concentration of the reactants, results in more reacting particles per unit volume. The particles are closer together and thus have a higher frequency of effective collisions. Rate of reaction will increase.

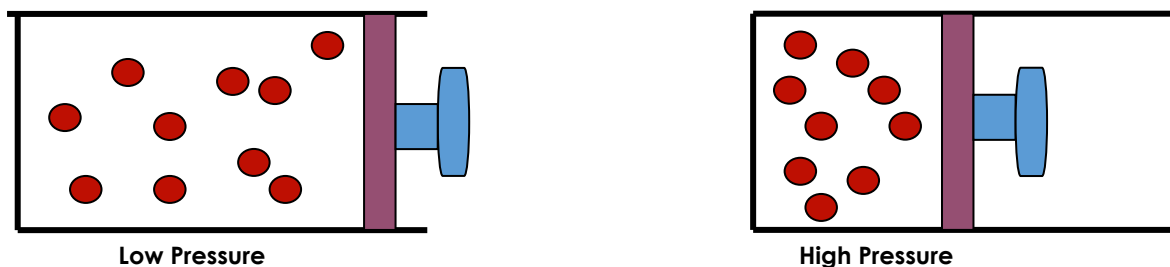
Hence the speed of reaction increases with higher concentration of reactants

Volume of Gas

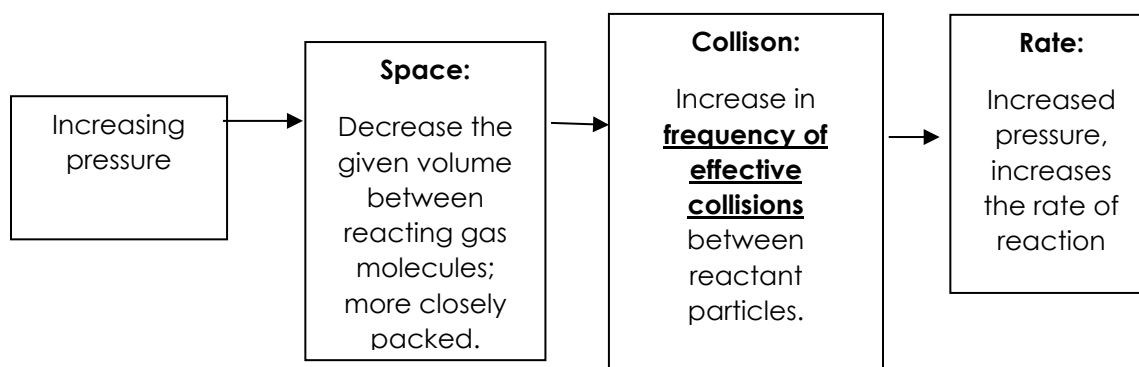




(b) Pressure of gaseous reactants



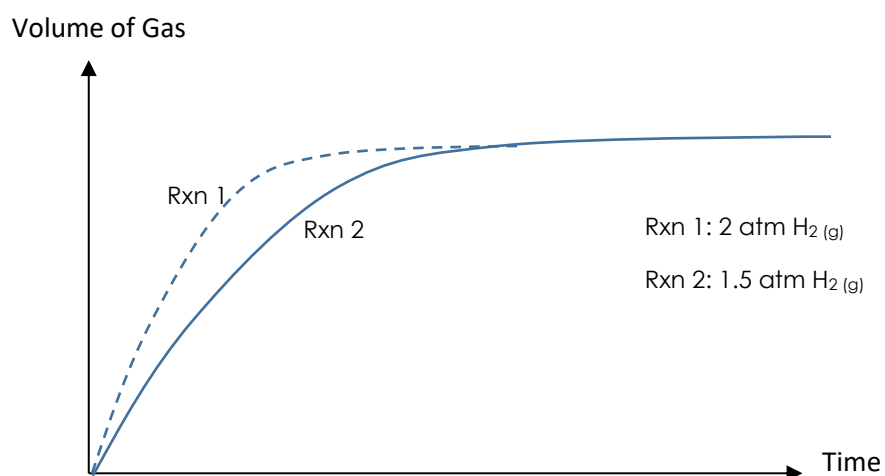
FLOW CHART EXPLANATION



EXPLANATION

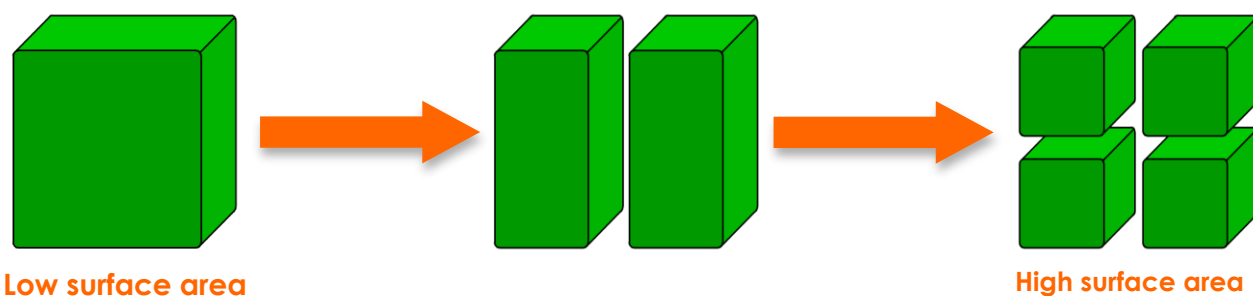
Increasing pressure in gaseous reactants, results in smaller volume for the same amount of reacting gas molecules. The reactant gas molecules are closer to each other, thus there will be higher frequency of effective collisions. Rate of reaction will increase.

Hence if the pressure for gaseous reactants is increased, the rate of reaction will be faster.

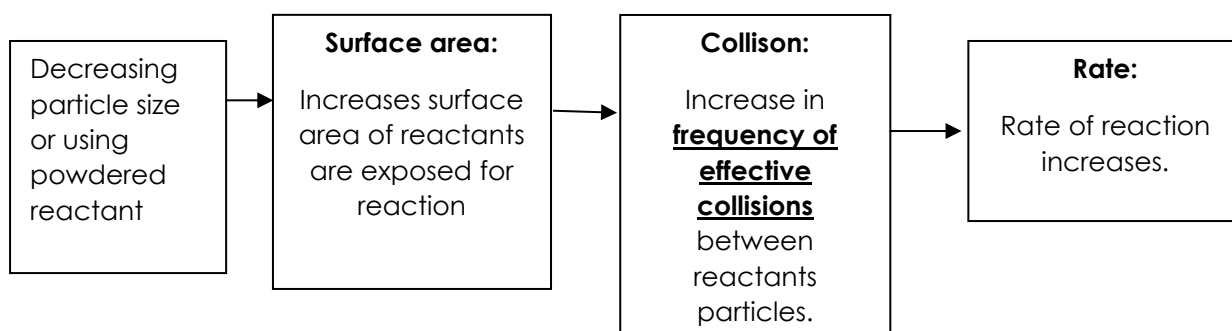




(c) Particle size or surface area of particles

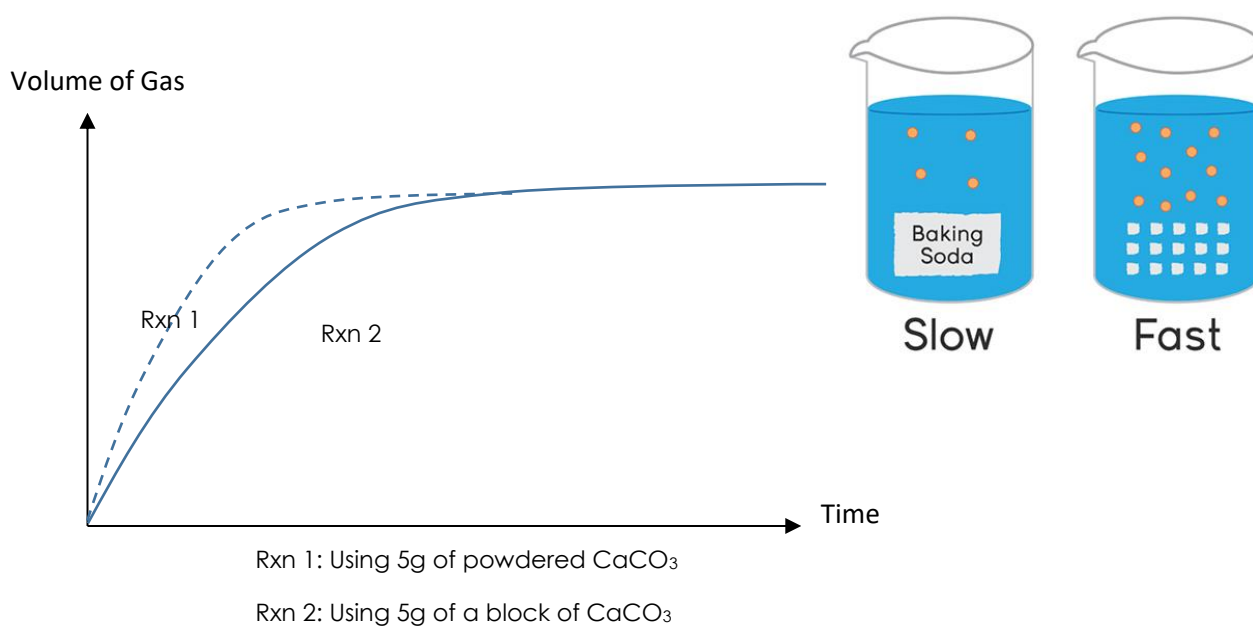


FLOW CHART EXPLANATION



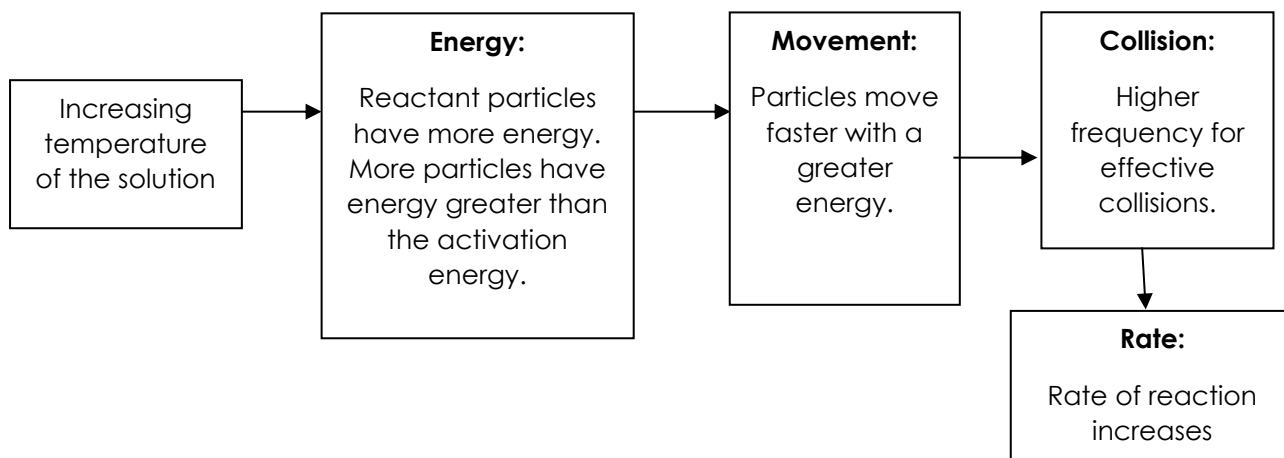
EXPLANATION

Decreasing the size of the particles results in higher total surface area of the reactants exposed for effective collisions. This results in a higher frequency of effective collisions. Hence increased surface area of reactants results in a faster rate of reaction.





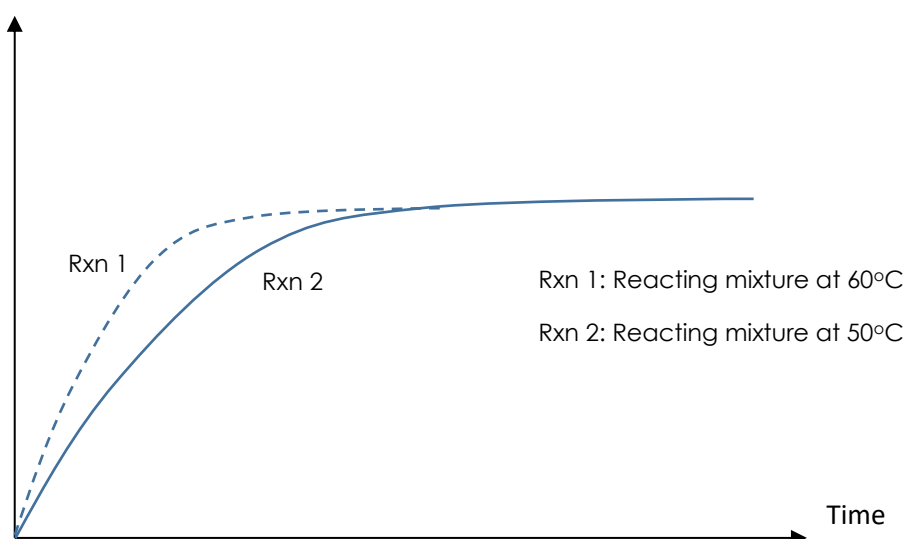
(d) Increasing temperature



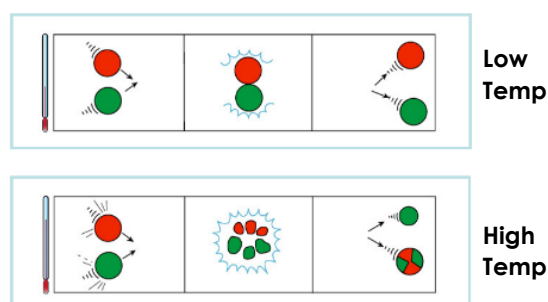
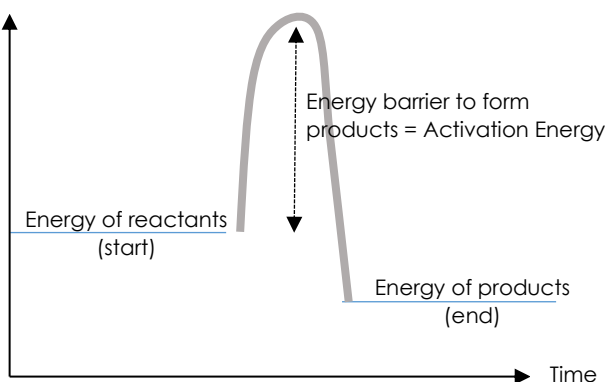
EXPLANATION

Increasing the temperature of a reaction results in reactant particles having more energy and the particles will move faster. More particles now have energy greater than the activation energy to form products. The frequency of effective collisions increases. The rate of reaction increases.

Volume of Gas



Energy



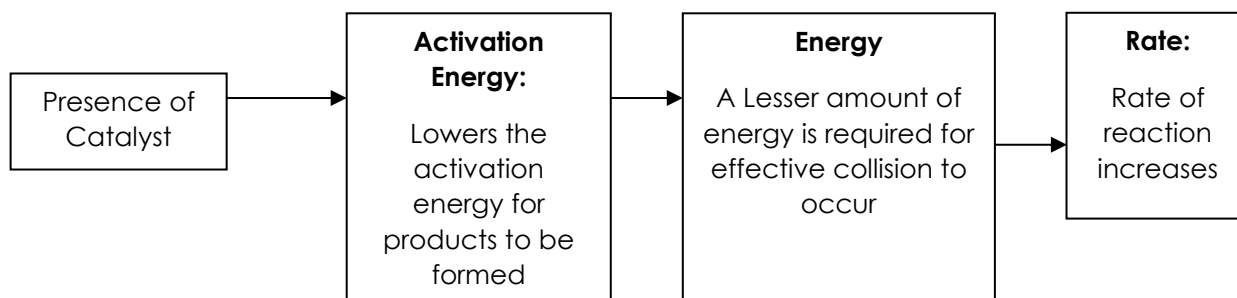


(e) Catalyst

Definition of a Catalyst:

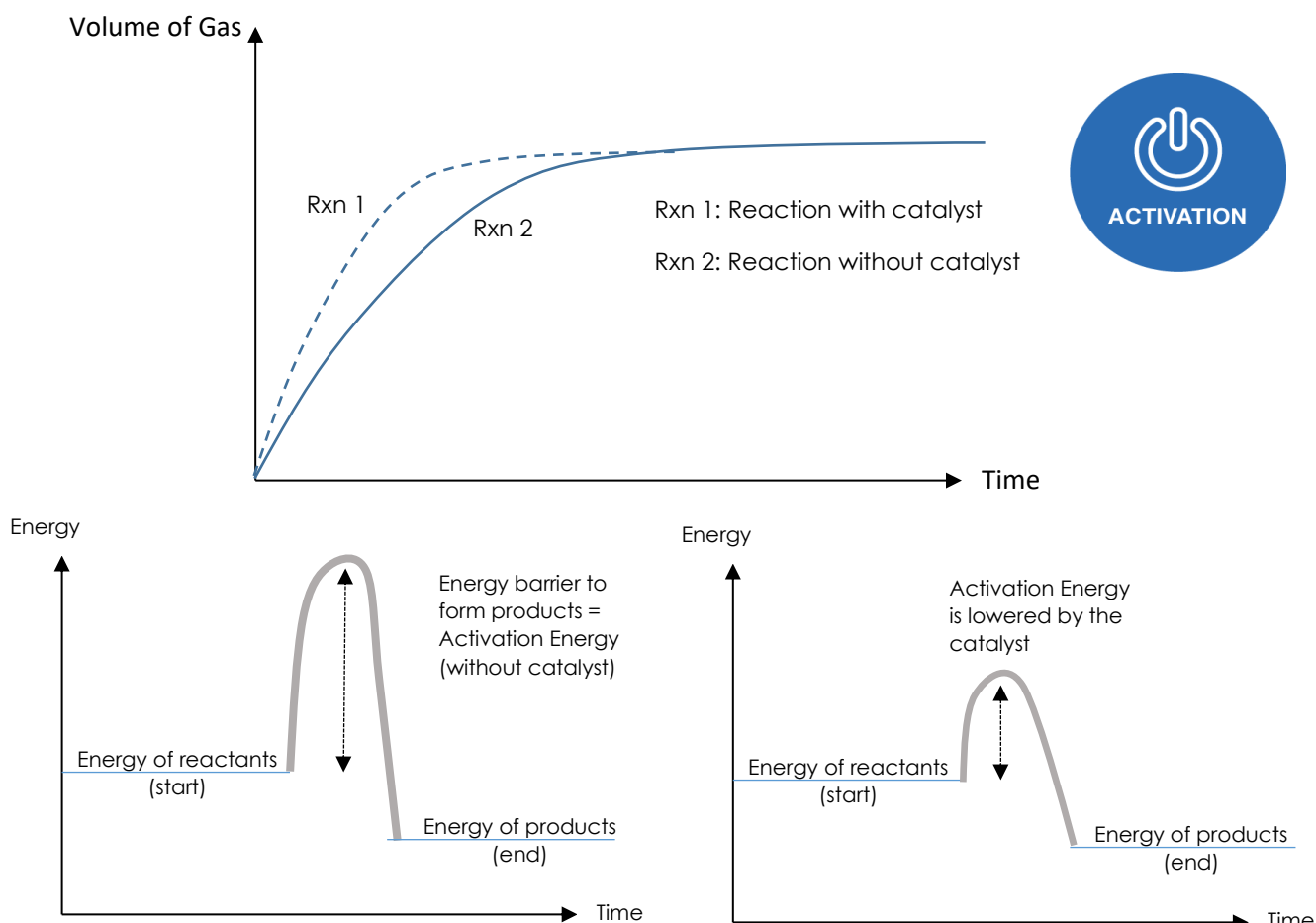
A catalyst is a substance which **increases the speed** of a chemical reaction and **remains chemically unchanged** at the end of the reaction. They provide alternative pathways that requires lower activation energy to form products.

Not all reaction can have a catalytic reaction. Catalyst are also specific in nature, they are specific in their reaction and cannot react the same for other reaction.



EXPLANATION

Presence of catalyst, decreases the activation energy required to form products. More reactants can now reach the lowered activation energy at a shorter time. Lesser amount of energy (lower temperature needed) is also required for effective collisions. The rate of reaction increases.





PART IV: EXAMPLES OF CATALYST

Catalyst	Reaction
Manganese (IV) oxide MnO_2	Used to speed up the decomposition of hydrogen peroxide (H_2O_2) to hydrogen and water
Iron (Pure Chemistry)	Used in the Haber process to form Ammonia gas
Vanadium (V) oxide V_2O_5 (Not in syllabus)	Used in the manufacture of sulfuric acid
Platinum or Rhodium	Used in Catalytic convertors
Aluminium oxide or Silicon oxide	Used in cracking of organic compounds to form hydrogen gas
Nickel	Used in the hydrogenation of alkenes to alkanes
Enzymes - Living Biological Catalyst (Yeast, Amylase)	<p>Biological catalyst found in plants and animals that speeds up chemical reaction.</p> <p>They are sensitive to temperature (Too high a temperature can <u>destroyed/denatured them</u>, too low a temperature can render them <u>inactive.</u>)</p> <p>They are also sensitive to pH.</p>





SUMMARY TABLE

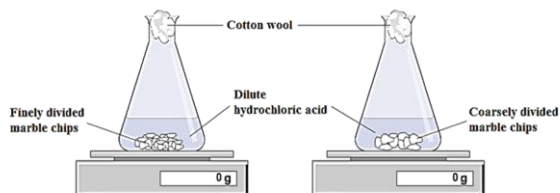
Factor	Effect on reacting particles		Effect of Rate
Concentration	Particles are closer together in a fixed given volume	Higher frequency of effective collisions	Rate of Reaction is increased
Pressure (Affects gas only)	Gaseous particles are forced closer together	Higher frequency of effective collisions between gaseous molecules	
Surface area	Increase surface area increases the contact surface area of reacting particles	More surface area for reacting is exposed for higher frequency of effective collisions.	
Temperature	Reacting particles have higher kinetic energy when temperature is increased. (Heat energy converted into KE of particles)	More particles thus have energy greater than the activation energy required, thus higher frequency of effective collisions.	
Catalyst	Activation energy for the reaction is lowered	More particles thus have energy greater than the activation energy required, thus higher frequency of effective collisions.	



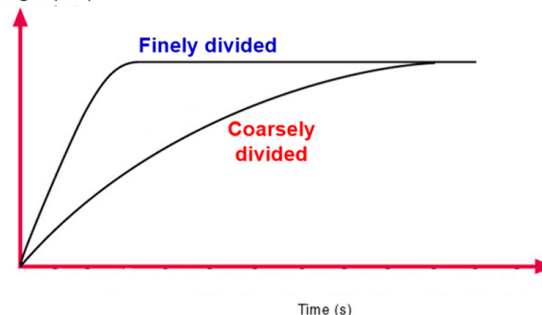
PART V: EXPERIMENTS SETUP

Effect of Particle Size

Marble chips of the same mass are added to equal volumes of hydrochloric acid (in excess) in 2 conical flasks as shown in the diagram above.

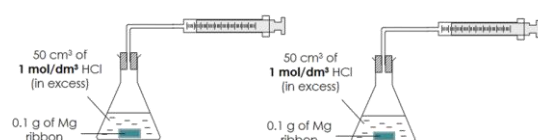


Volume of carbon dioxide gas (cm³)



Effect of Temperature

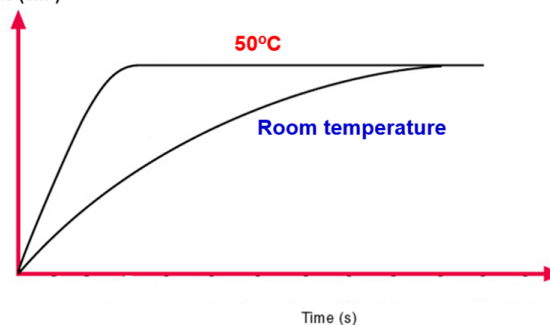
Carry out the reactions below in a beaker and measure the volume of gas given off at regular time interval using the gas syringe. Plot the graph of volume of gas collected against time for both reactions A and B.



Reaction A
At room temperature

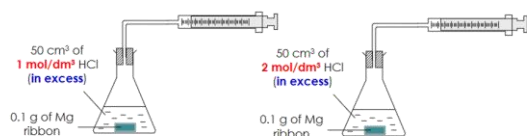
Reaction B
Placed in a water bath at 50°C

Volume of hydrogen gas (cm³)



Effect of Concentration

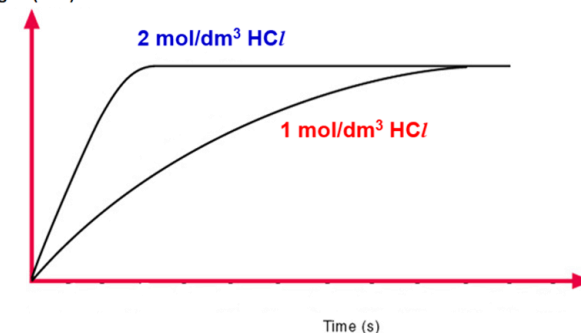
Carry out the reactions below in a beaker and measure the volume of gas given off at regular time interval using the gas syringe. Plot the graph of volume of gas collected against time for both reactions A and B.



Reaction A

Reaction B

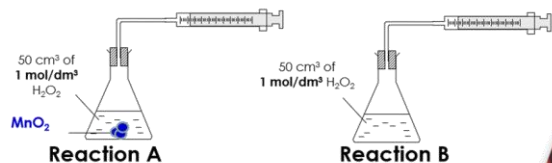
Volume of hydrogen gas (cm³)



Effect of catalysts



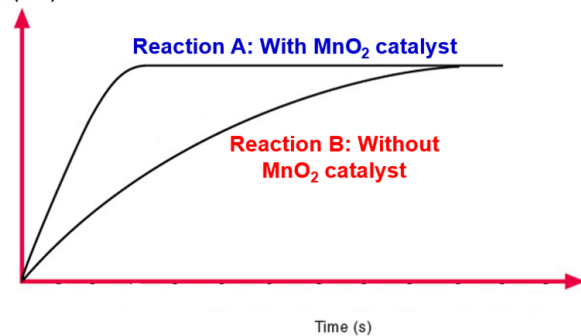
2 conical flasks, each containing the same volume of hydrogen peroxide with the same concentration are left on a table. One of the test tubes has **manganese oxide(IV)** added, the second one is not catalysed. The volume of gas given off at regular time interval was measured according to the diagram below.



Reaction A

Reaction B

Volume of oxygen gas (cm³)





PART VI: GRAPH INTERPRETATION AND APPLICATION

In a chemical reaction,



Assuming that Mg is the excess reactant and HCl is the limiting reactant, the number of moles of HCl used will determine the moles of H₂ gas produced.

Hence from the table

Mole of HCl used	Mole of H ₂ produced
6 moles	3 moles
10 moles	5 moles

Application:

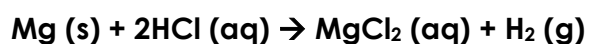
Assume HCl is the limiting reagent, the concentration and volume of the acid used will affect the number of HCl used and consequently the moles of H₂ evolved.

Con of HCl (mol/dm ³)	Vol of HCl (cm ³)	Mol of HCl	Mol of H ₂
1.0	25	0.025	0.0125
2.0	25	0.05	0.025
2.0	50	0.1	0.05

Qns 1:

Sketch the graph for the following reactions on the same axes.

Assume Mg is the excess reactant



Reaction 1: 1.0 mol/dm³ and 25 cm³ of HCl used.

Reaction 2: 2.0 mol/dm³ and 25 cm³ of HCl used.

Volume of Gas



Step 1: Work out the moles of

HCl for reaction 1 & 2

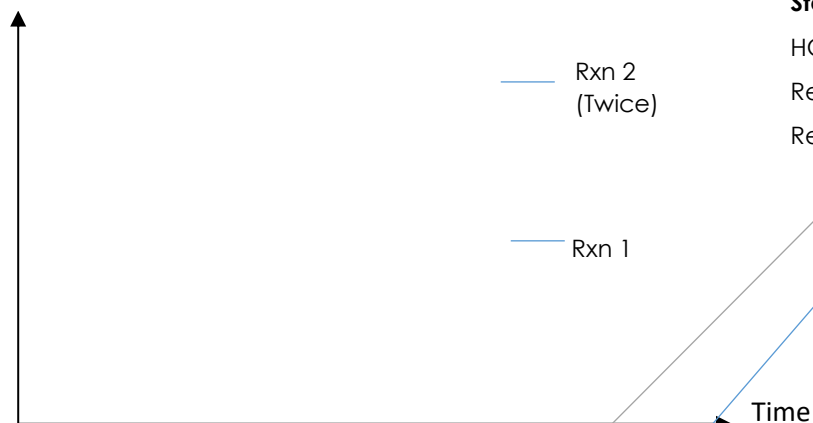
Reaction 1: 0.025 mol

Reaction 2: 0.05 mol (twice)

Graph volume of gas must be proportional



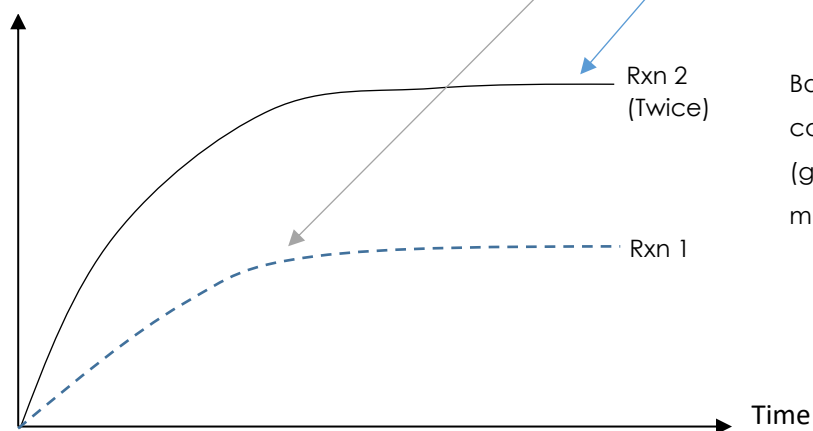
Volume of Gas

**Step 2:** Look at the concentration of

HCl for reaction 1 & 2

Reaction 1: 1.0 mol/dm³Reaction 2: 2.0 mol/dm³

Volume of Gas

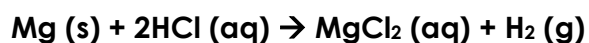


Based on collision theory, a higher concentration will mean a faster rate (gradient) Hence grad for reaction 2 must be steeper than reaction 1

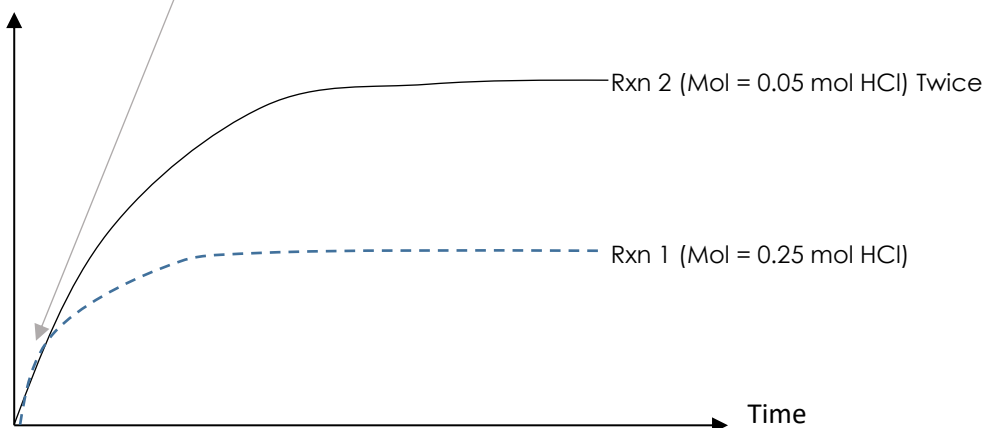
Qns 2:

Sketch the graph for the following reactions on the same axes.

Assume Mg is the excess reactant

Reaction 1: 1.0 mol/dm³ and 25 cm³ of HCl used.Reaction 2: 1.0 mol/dm³ and 50 cm³ of HCl used.

Volume of Gas



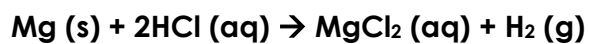
The gradient (rate) is the same since both HCl has concentration of 1.0 mol/dm³



Qns 3:

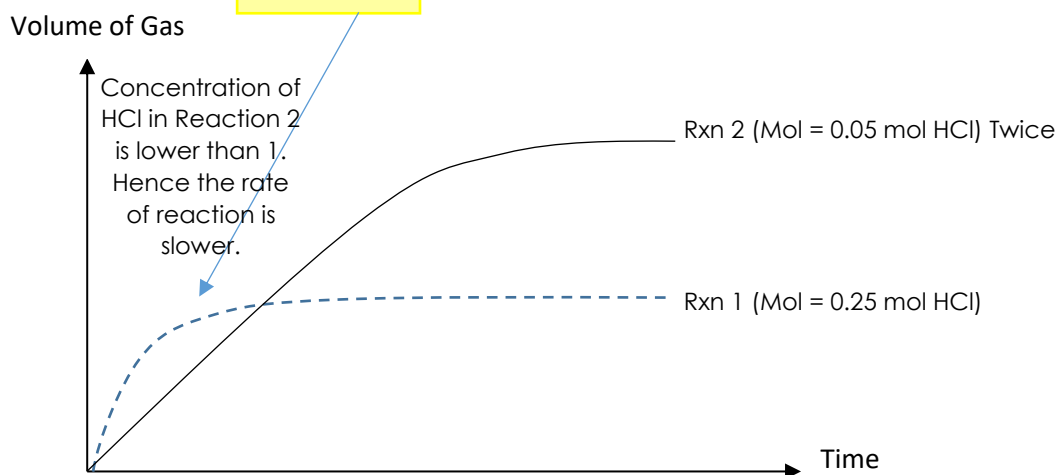
Sketch the graph for the following reactions on the same axes.

Assume Mg is the excess reactant



Reaction 1: 1.0 mol/dm³ and 25 cm³ of HCl used.

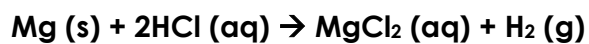
Reaction 2: 0.5 mol/dm³ and 100 cm³ of HCl used.



Qns 4

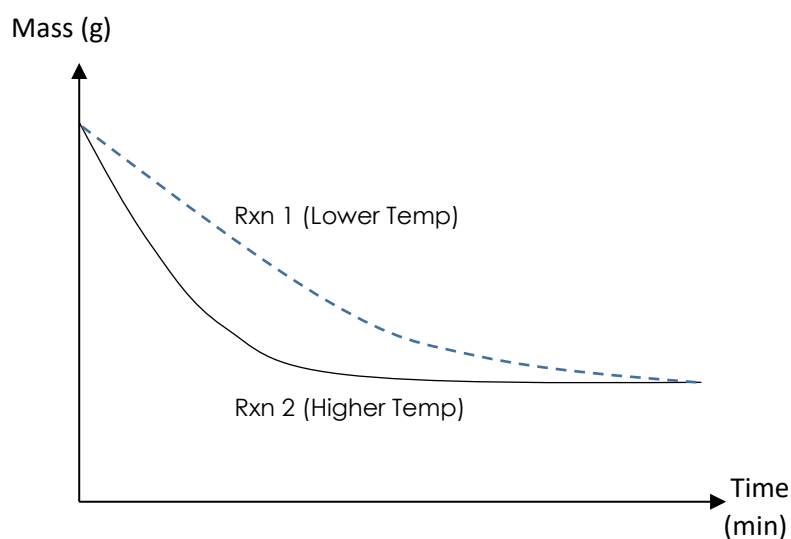
Sketch the graph for the following reactions on the same axes.

Assume Mg is the excess reactant



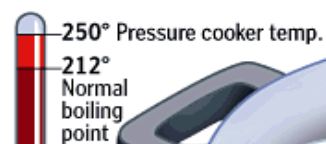
Reaction 1: Temperature of mixture 40°C

Reaction 2: Temperature of mixture 70°C



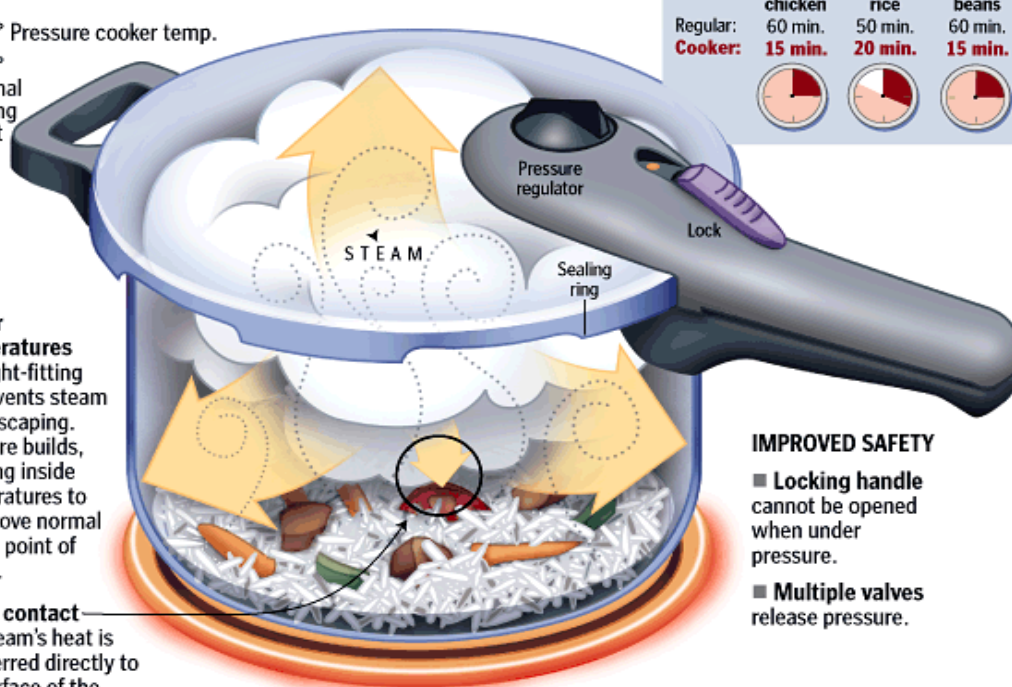


APPLICATIONS



1 Higher temperatures
The tight-fitting lid prevents steam from escaping. Pressure builds, allowing inside temperatures to rise above normal boiling point of 212°F.

2 Direct contact
The steam's heat is transferred directly to the surface of the food.

COOKING TIMES
(approximate)

	Whole chicken	Brown rice	Black beans
Regular:	60 min.	50 min.	60 min.
Cooker:	15 min.	20 min.	15 min.

IMPROVED SAFETY

- Locking handle cannot be opened when under pressure.
- Multiple valves release pressure.

Food gets cooked faster due to higher boiling point of water in pressure



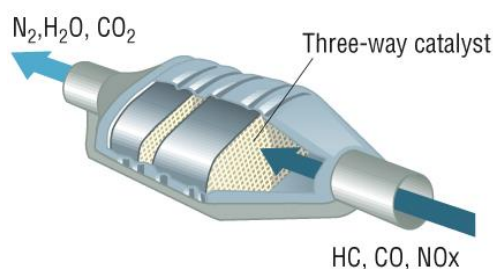
Food remains fresh for a longer period of time when they are kept in a freezer than left outside. The **lower temperature** inside freezer **decreases the rate** of food spoilage process.

Systemic enzymes have been shown to break down the outer fibrin wall of the cancer cell making it easier for chemotherapy to penetrate and destroy the cancer cells

HIGH-ENZYME FOODS

- Avocado
- Papaya
- Pineapple
- Aloe Vera
- Barley grass
- Cucumbers
- Garlic
- Ginger root
- Pau d'arco
- Sprouts
- Wheatgrass juice
- Coconut
- Flaxseed
- Chlorella
- Spirulina
- Figs

rawforbeauty.com



Platinum is a catalyst in the **catalytic converters** of car exhausts. It catalyzes the conversion of carbon monoxide and nitrogen oxide into the less polluting carbon dioxide and nitrogen.



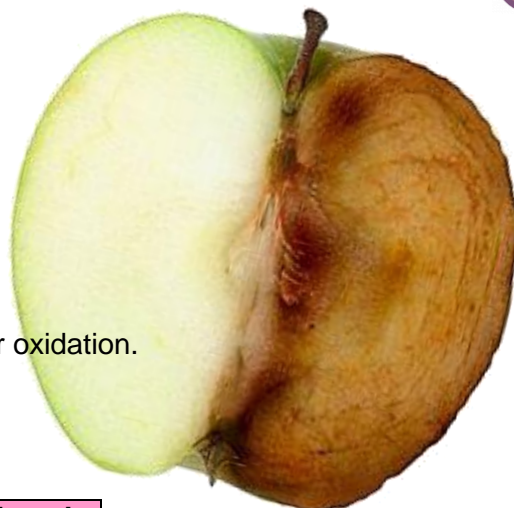
Redox Rxns

Method involves the following reaction:

Redox reactions consists of:

- (a) **Oxidation** and
- (b) **Reduction** where both occur simultaneously.

*There are some reactions that do not have reduction nor oxidation.



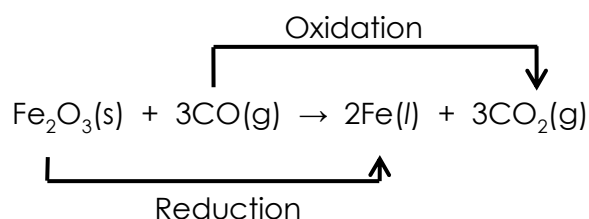
Definition of Oxidation and Reduction

A substance is oxidised when it	A substance is reduced when it
Gains oxygen	Loses oxygen
Loses hydrogen	Gains hydrogen
Loses electrons	Gains electrons
Increase in oxidation state	Decrease in oxidation state

Part I: Oxidation & Reduction in terms of Oxygen

Oxidation - Gain of Oxygen	1. $2\text{Ca(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{CaO(s)}$ • Ca has gained oxygen to form CaO. • Ca has been oxidised to CaO.
	2. $2\text{Mg(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{MgO(s)}$ • Mg has gained oxygen to form MgO. • Mg has been oxidised to MgO.
Reduction — Loss of Oxygen	1. $\text{Zn(s)} + \text{CuO(s)} \rightarrow \text{ZnO(s)} + \text{Cu(s)}$ • CuO has lost oxygen to form Cu. • CuO has been reduced to Cu.
	2. $\text{Fe}_2\text{O}_3\text{(s)} + 3\text{CO(g)} \rightarrow 2\text{Fe(l)} + 3\text{CO}_2\text{(g)}$ • Fe_2O_3 has lost oxygen to form Fe. • Fe_2O_3 has been reduced to Fe.

Redox Reaction shown below occurs, when oxygen is gained and lost at the same time



Oxidation — CO has **gained oxygen** to form CO_2 .

Reduction — Fe_2O_3 has **lost oxygen** to form Fe.



Part II: Oxidation & Reduction in terms of Hydrogen

Oxidation - Loss of Hydrogen	1. $\text{H}_2\text{S}(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g}) + \text{S}(\text{s})$ <ul style="list-style-type: none"> H_2S has lost hydrogen to form S. H_2S has been oxidised to S.
	2. $2\text{NH}_3(\text{g}) + 3\text{CuO}(\text{s}) \rightarrow \text{N}_2(\text{g}) + 3\text{Cu}(\text{s}) + 3\text{H}_2\text{O}(\text{g})$ <ul style="list-style-type: none"> NH_3 has lost hydrogen to form N_2. NH_3 has been oxidised to N_2.
Reduction — Gain of Hydrogen	1. $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g})$ <ul style="list-style-type: none"> Cl_2 has gained hydrogen to form HCl. Cl_2 has been reduced to HCl.
	2. $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$ <ul style="list-style-type: none"> N_2 has gained hydrogen to form NH_3. N_2 has been reduced to NH_3.

Part III: Oxidation & Reduction in terms of Electrons

Oxidation - Loss of Electrons	1. $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$ <ul style="list-style-type: none"> Mg has lost two electrons to form Mg^{2+}. Mg has been oxidised to Mg^{2+}.
	2. $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ <ul style="list-style-type: none"> Fe^{2+} has lost one electron to form Fe^{3+}. Fe^{2+} has been oxidised to Fe^{3+}.
Reduction — Gain of Electrons	1. $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ <ul style="list-style-type: none"> Cu^{2+} has gained two electrons to form Cu. Cu^{2+} has been reduced to Cu.
	2. $\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$ <ul style="list-style-type: none"> Fe^{3+} has gained one electron to form Fe^{2+}. Fe^{3+} has been reduced to Fe^{2+}.

Redox Reaction shown below occurs, when electrons is gained and lost at the same time

Example 1: $\text{Cu}(\text{s}) + 2\text{AgNO}_3(\text{aq}) \rightarrow \text{Cu}(\text{NO}_3)_2(\text{aq}) + 2\text{Ag}(\text{s})$ (Overall equation)

Ionic equation: $\text{Cu}(\text{s}) + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{Ag}(\text{s})$

- Cu has lost two electrons to form Cu^{2+} .
- Cu has been oxidised to Cu^{2+} .
- Ag^+ has gained an electron to form Ag.
- Ag^+ has been reduced to Ag.



Part III: Oxidation & Reduction in terms of Oxidation number/ state

- Oxidation state is the charge an atom of an element would have if it existed as an ion in a compound.
- When a substance loses or gains electrons, its oxidation state changes.

Rules	Example	Oxidation state
The oxidation state of a pure element is ZERO	Cu	0
	S	0
	Cl ₂	0
The oxidation state of a simple ion is the same as the charge of the ion.	K ⁺	+1
	Zn ²⁺	+2
	Cl ⁻	-1
	O ²⁻	-2
The oxidation states of the atoms present in the formula of a compound must add to 0.	In MgCl ₂ ,	
	Sum of oxidation numbers = +2 + 2(-1) = 0	
	Mg	+2
	Cl	-1 (x2 for Cl atoms)
	In CaCO ₃ ,	
	Sum of oxidation numbers = +2 + 4 + 3(-2) = 0	
	Ca	+2
The total of the oxidation states of the atoms in a polyatomic ion is equal to the overall charge on the ion.	C	+4
	O	-2 (x 3 for O atoms)
	In SO ₄ ²⁻	
	S	+6
	O	-2 (x 4 for O atoms)

Exceptional Cases

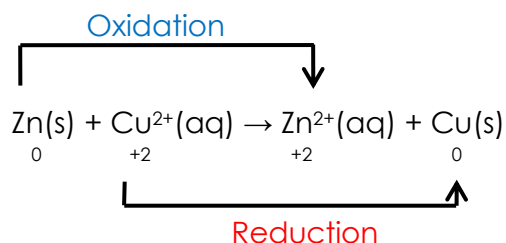
In Hydrogen peroxide, O has an oxidation of -1	In H ₂ O ₂	
	H	+1
	O	-1
In metal hydride, the more positive metal will retain its positive charge, whereas H will be -1	For CaH ₂	
	Ca	+2
	H	-1



Oxidation and Reduction in Terms of Changes in Oxidation State

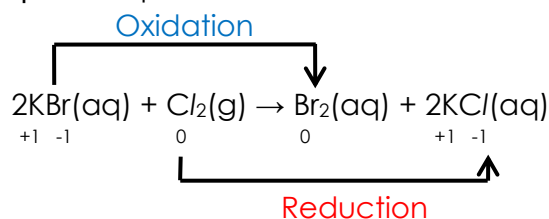
- Oxidation occurs when the oxidation state of a substance increases.
- Reduction occurs when the oxidation state of a substance decreases.

Example 1: Displacement reaction between metals

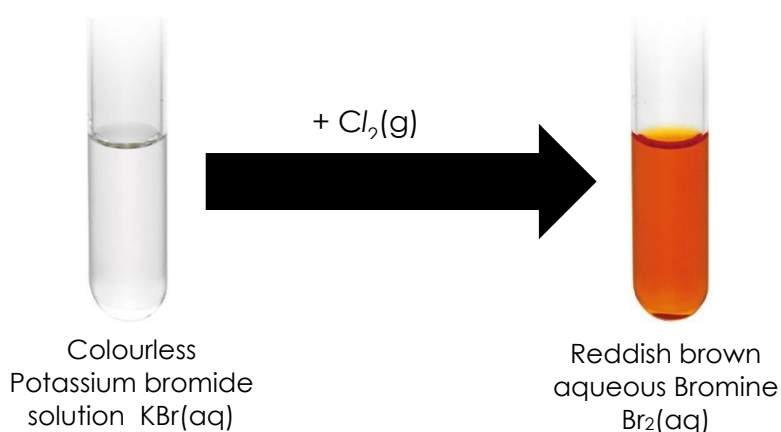


- The oxidation state of zinc has increased from 0 (in Zn) to +2 (in Zn^{2+}).
- Zn has been oxidised to Zn^{2+} .
- The oxidation state of copper has decreased from +2 (in Cu^{2+}) to 0 (in Cu).
- Cu^{2+} has been reduced to Cu.

Example 2: Displacement reaction between halogens



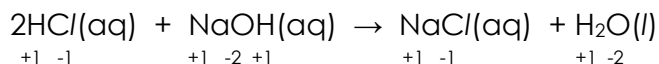
- The oxidation state of bromine has increased from -1 (in KBr) to 0 (in Br_2).
- KBr has been oxidised to Br_2 .
- The oxidation state of chlorine has decreased from 0 (in Cl_2) to -1 (in KCl).
- Cl_2 has been reduced to KCl.





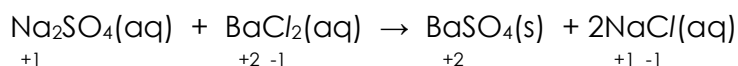
Part IV: Non-Redox Reactions

1. Neutralisation reactions (Acid + Metal hydroxide)



No change in oxidation states of hydrogen and oxygen. Hence, neutralisation reactions are not redox reactions.

2. Precipitation reactions (Insoluble salts are formed)

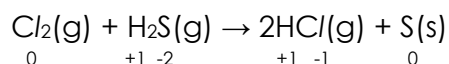


The oxidation of sulfate ions need not be calculated as the formula remains the same on both side of the equation. No change in oxidation states of barium, sulfur and oxygen. Hence, precipitation reactions are not redox reactions.

Part V: Oxidizing and Reducing agents

Oxidizing agent	Reducing agent
<ul style="list-style-type: none"> causes another substance to be oxidised by <ul style="list-style-type: none"> a) giving oxygen; b) removing hydrogen; c) accepting electrons; 	<ul style="list-style-type: none"> causes another substance to be reduced by <ul style="list-style-type: none"> a) removing oxygen; b) giving hydrogen; c) donating electrons.
<ul style="list-style-type: none"> is reduced at the end of the reaction. 	<ul style="list-style-type: none"> is oxidised at the end of the reaction.

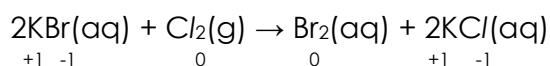
Example 1:



<ul style="list-style-type: none"> Cl_2 has been reduced. The oxidation of chlorine has decreased from 0 (in Cl_2) to -1 (in HCl) 	<ul style="list-style-type: none"> H_2S has been oxidised. The oxidation of sulfur has increased from -2 (in H_2S) to 0 (in S)
Thus, Cl_2 is an oxidising agent.	Thus, H_2S is a reducing agent.

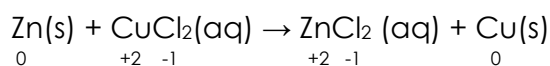


Example 2:



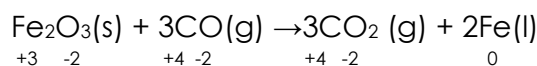
<ul style="list-style-type: none"> Cl_2 has been reduced to KCl. The oxidation of chlorine has decreased from 0 (in Cl_2) to -1 (in KCl) 	<ul style="list-style-type: none"> KBr has been oxidized to Br_2. The oxidation of bromine has increased from -1 (in KBr) to 0 (in Br_2)
Thus, Cl_2 is an oxidising agent.	Thus, KBr is a reducing agent.

Example 3:

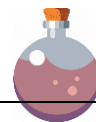


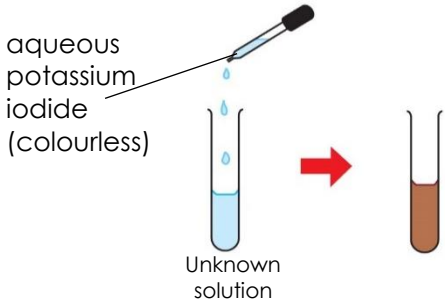
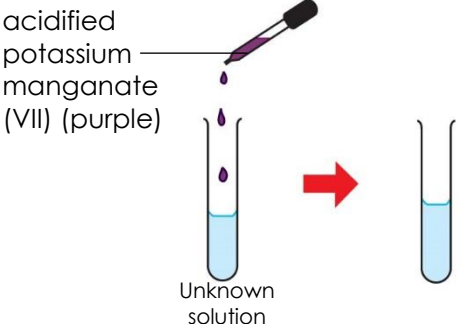
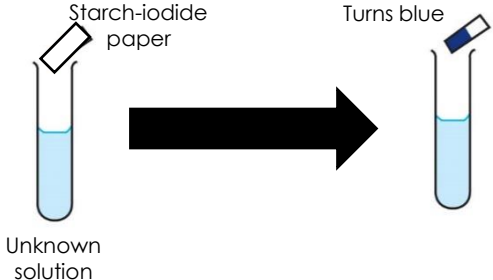
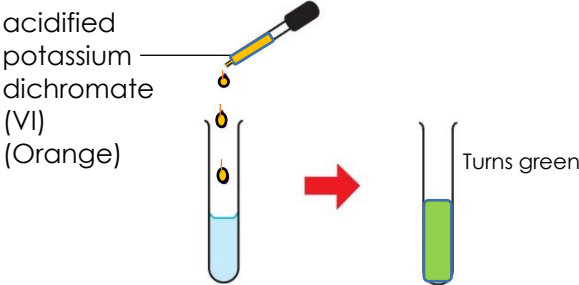
<ul style="list-style-type: none"> CuCl_2 has been reduced to Cu. The oxidation of copper has decreased from +2 (in CuCl_2) to 0 (in Cu) 	<ul style="list-style-type: none"> Zn has been oxidized to ZnCl_2. The oxidation of Zinc has increased from 0 (in Zn) to +2 (in ZnCl_2)
Thus, CuCl_2 is an oxidising agent.	Thus, Zn is a reducing agent.

Example 4:



<ul style="list-style-type: none"> Fe_2O_3 has been reduced to Fe. The oxidation of iron has decreased from +3 (in Fe_2O_3) to 0 (in Fe) 	<ul style="list-style-type: none"> CO has been oxidized to CO_2. The oxidation of carbon has increased from +2 (in CO) to +4 (in CO_2)
Thus, Fe_2O_3 is an oxidising agent.	Thus, CO is a reducing agent.



Test for	
The presence of Oxidizing agents	The presence of Reducing agents
<p>Test 1: Add aqueous potassium iodide, KI, to the unknown solution.</p>  <p>If the resultant mixture turns brown. We can conclude that the unknown solution contains an oxidizing agent.</p> <p>Explanation: Iodide ions are oxidised to iodine by the oxidising agent. The brown solution indicates the presence of I₂</p> $2\text{I}^-(\text{aq}) \rightarrow \text{I}_2(\text{aq}) + 2\text{e}^-$	<p>Test 1: Add acidified potassium manganate(VII), KMnO₄, to the unknown solution.</p>  <p>If the resultant mixture turns colourless. We can conclude that the unknown solution contains a reducing agent.</p> <p>Acidified potassium manganate(VII) solution will turn from purple to colourless when reduced.</p> <p>Explanation for using KMnO₄: Manganate(VII) ion MnO₄⁻ is reduced to manganese(II) ion Mn²⁺.</p> $\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$ <p>Oxidation state of Mn decreases from +7 to +2.</p>
<p>Test 2: Dip a piece of starch-iodide paper in the unknown solution.</p>  <p>If the paper from white to blue. We can conclude that the unknown solution contains an oxidizing agent.</p> <p>Explanation:</p> <ul style="list-style-type: none"> Iodide ions are oxidised to iodine by the oxidising agent. The iodine then reacts with the starch to give a blue colour. 	<p>Test 2: Add acidified potassium dichromate (VI), K₂Cr₂O₇, to the unknown solution.</p>  <p>If the resultant mixture turns green. We can conclude that the unknown solution contains a reducing agent.</p> <p>Acidified potassium dichromate(VI) solution will turn from orange to green when reduced.</p> $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+ + 6\text{e}^-(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$ <p>Oxidation state of Cr decreases from +6 to +3.</p>



Examples of	
Oxidizing agent	Reducing agent
<ul style="list-style-type: none">• Bromine (Br_2)• Chlorine (Cl_2)• Oxygen (O_2)• Hydrogen peroxide (H_2O_2)• Potassium manganate(VII) (KMnO_4) \rightarrow purple turns colourless if RA present• Potassium dichromate(VI) ($\text{K}_2\text{Cr}_2\text{O}_7$) \rightarrow orange turns green if RA present• Concentrated sulfuric acid (H_2SO_4)• Nitric acid (HNO_3)	<ul style="list-style-type: none">• Metals• Carbon (C)• Carbon monoxide (CO)• Hydrogen (H_2)• Hydrogen sulfide (H_2S)• Sulfur dioxide (SO_2)• Ammonia (NH_3)• Potassium iodide (KI) \rightarrow colorless turns brown if OA present

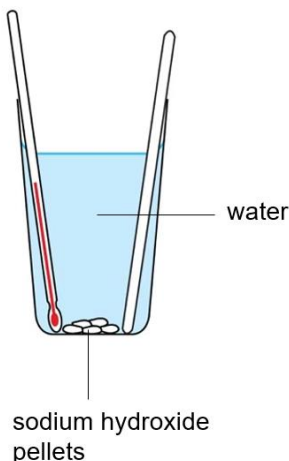
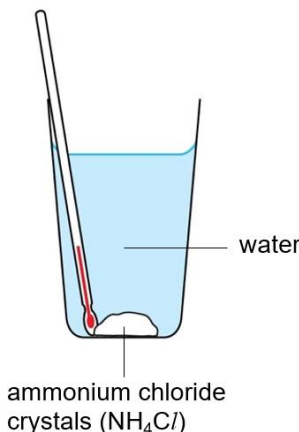


Energy in Chemistry

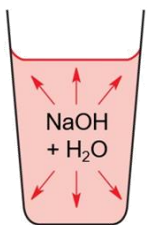
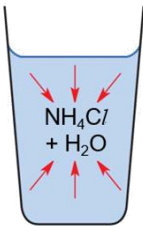
Energy Changes

- Energy cannot be created or destroyed, but it can be changed from one form to another.
- All chemical reactions involve energy changes.
- An energy change is sometimes observed as a rise or drop in temperature.

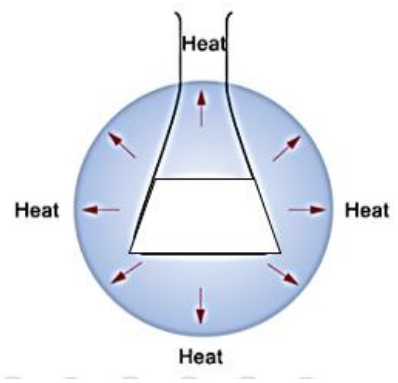
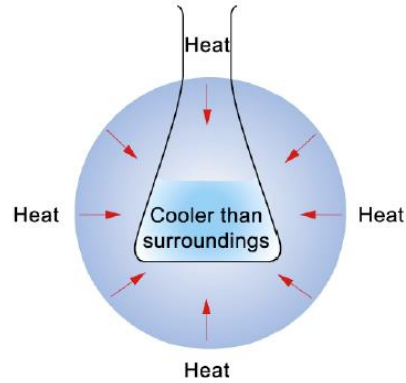
Experiments to test for the Exothermic and Endothermic changes

Exothermic changes	Endothermic changes												
Test 1 1. Measure the temperature of the water. 2. Add sodium hydroxide pellets and stir to dissolve. 3. Record the highest temperature reached by the solution.	Test 2 1. Measure the temperature of the water. 2. Add solid ammonium chloride crystals and stir to dissolve. 3. Record the lowest temperature reached by the solution.												
													
Results													
<table><tr><th>Test</th><th>Solid</th><th>Initial temperature/°C</th><th>Final temperature/°C</th></tr><tr><td>1</td><td>NaOH</td><td>28</td><td>34</td></tr><tr><td>2</td><td>NH₄Cl</td><td>28</td><td>22</td></tr></table>	Test	Solid	Initial temperature/°C	Final temperature/°C	1	NaOH	28	34	2	NH ₄ Cl	28	22	<div>rise in temperature</div> <div>drop in temperature</div>
Test	Solid	Initial temperature/°C	Final temperature/°C										
1	NaOH	28	34										
2	NH ₄ Cl	28	22										



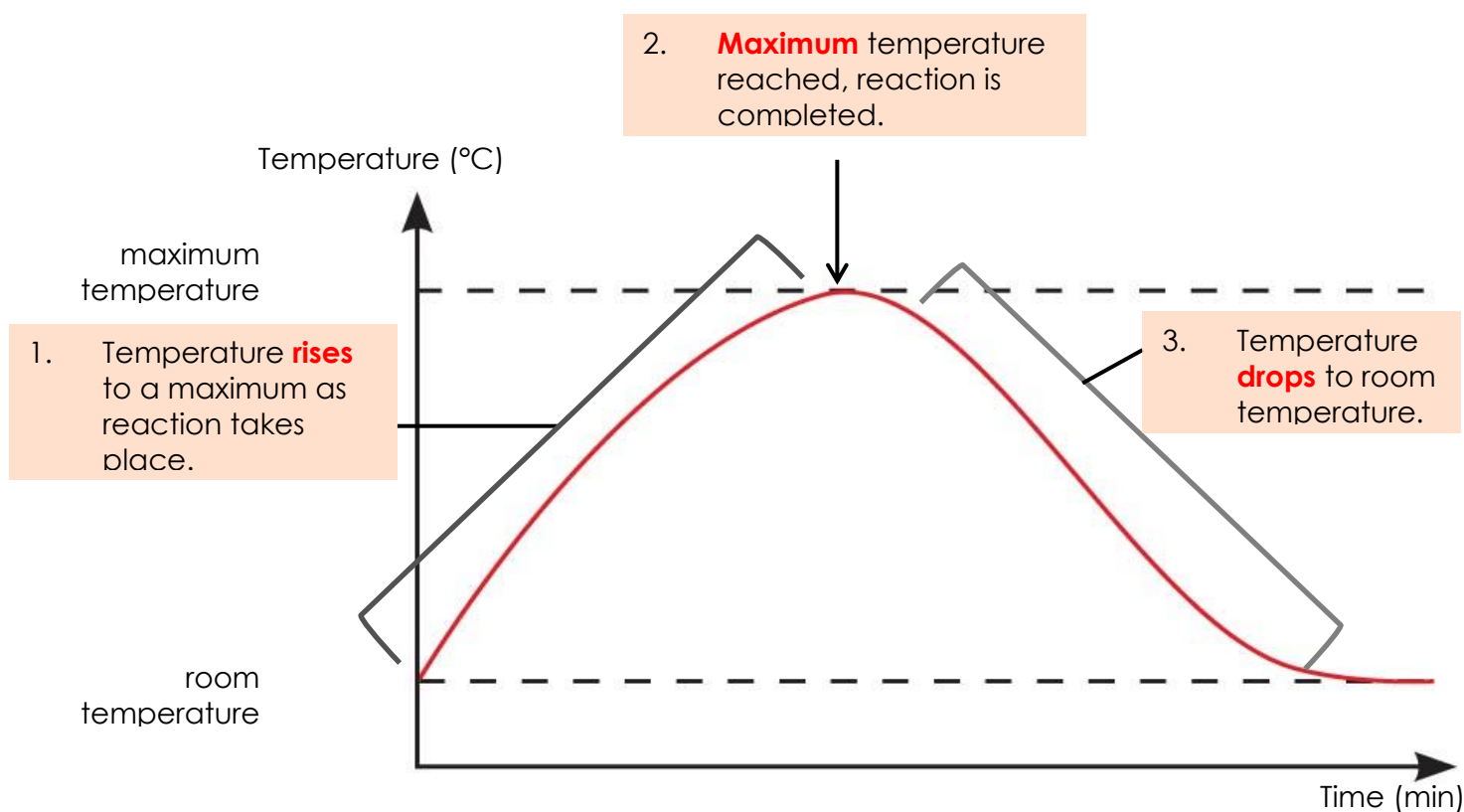
	
<p>Test 1 An exothermic change has taken place. Heat was given out <u>by reacting particles</u> to the surroundings.</p>	<p>Test 2 An endothermic change has taken place. Heat was absorbed <u>by reacting particles</u> from the surroundings.</p>

Definition of Exothermic and Endothermic Reactions

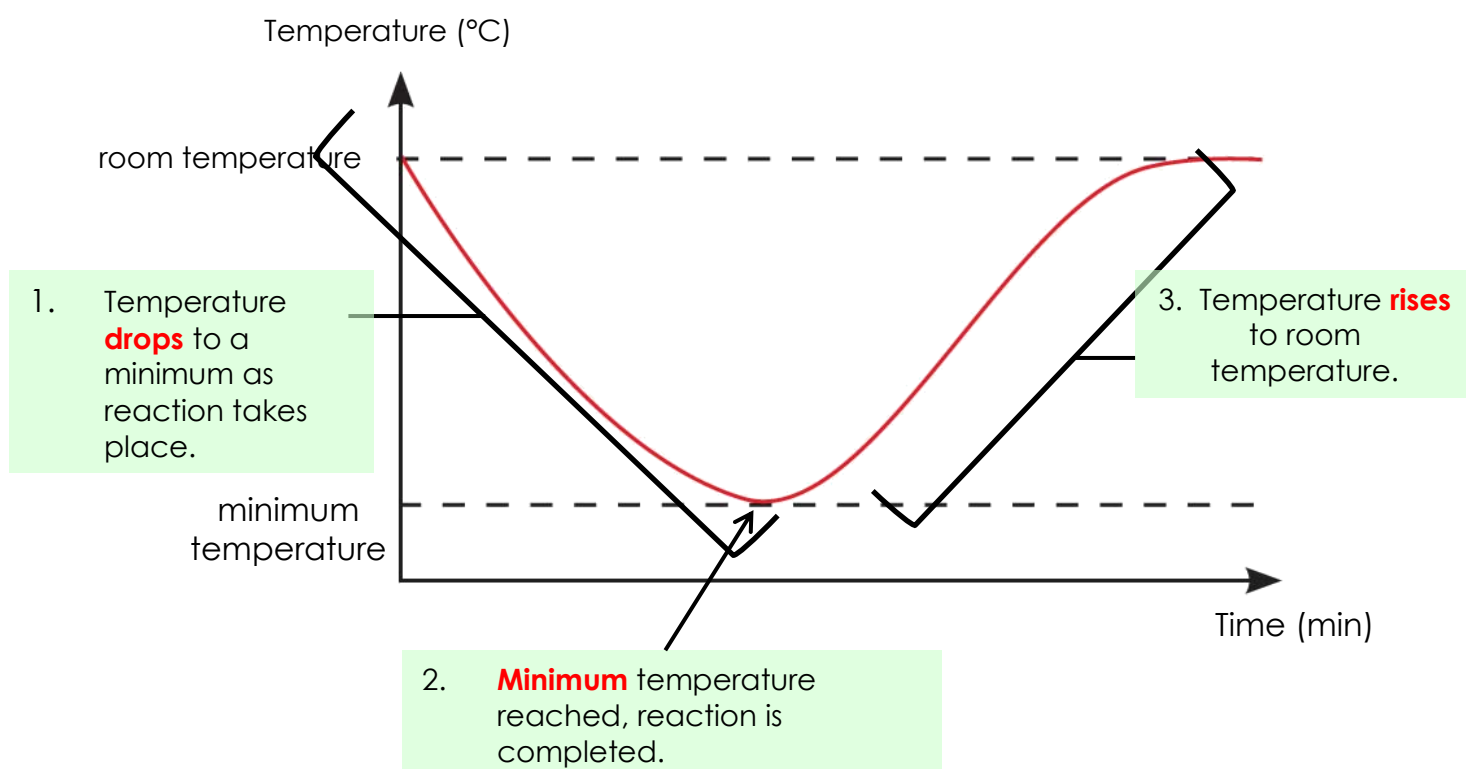
Exothermic Reaction	Endothermic Reaction
Reactions that give out heat energy to the surroundings are called exothermic reactions .	Reactions that absorb heat from the surroundings are called endothermic reactions .
Heat is transferred from the reactants to the surroundings.	Heat is transferred from the surroundings to the reactants.
Temperature of reaction mixture rises . Container feels hot . 	Temperature of reaction mixture falls. Container feels cooler. 



Graphs Showing Temperature Change for an Exothermic Reaction



Graphs Showing Temperature Change for a Endothermic Reaction





Examples of Exothermic and Endothermic Reactions

Exothermic Reaction	Endothermic Reaction
<ul style="list-style-type: none"> • Combustion • Neutralisation • Extraction of iron in the Blast Furnace • Corrosion of metals (e.g. rusting of iron) • Respiration • Condensation 	<ul style="list-style-type: none"> • Dissolving some ionic compounds (e.g. ammonium chloride, sodium carbonate crystals) in water • Thermal decomposition $\text{CaCO}_3 \xrightarrow{\text{heat}} \text{CaO} + \text{CO}_2$ • Photosynthesis - chlorophyll and other plant pigments <u>absorb energy (light) from the Sun</u> to convert carbon dioxide and water into glucose and oxygen. • Action of light on silver bromide in photographic film • Evaporation

Changes in Physical state

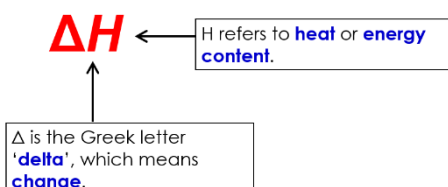
Melting (Solid → Liquid)	Freezing (Liquid → Solid)
<ul style="list-style-type: none"> • During the melting of ice, Heat energy is absorbed by the solid particles, causing them to vibrate faster and eventually overcome the forces of attraction between particles to form a liquid. • Heat energy is absorbed from the surroundings. • Melting is endothermic. 	<ul style="list-style-type: none"> • Particles lose heat energy to the surroundings, they vibrate slower and forces of attraction between particles increase. • Heat energy is lost to the surroundings. • Freezing is exothermic



Heat changes in Reaction

The amount of energy involved in a reaction is known as the **heat change** or **enthalpy change** of the reaction.

- measured in **kilojoules (kJ)**
- represented by the symbol **ΔH**



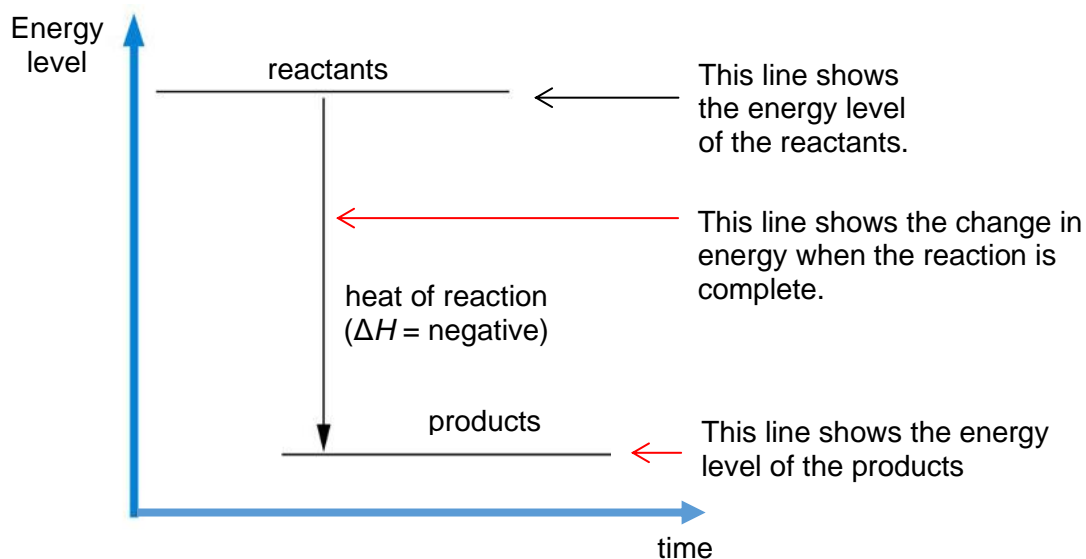
The **difference** between the energy levels of the products and the reactants is equal to the amount of energy given out or taken in (**ΔH**) by the reaction.

$$\Delta H = \text{total energy of Products} - \text{total energy of Reactants}$$

- For an **exothermic** reaction, **ΔH is negative (<0)** because the chemicals have **lost energy** to the surroundings.
- For an **endothermic** reaction, **ΔH is positive (>0)** because the chemicals have **gained energy** from the surroundings.

Energy Level Diagrams

An **energy level diagram** shows the energy change in a reaction.

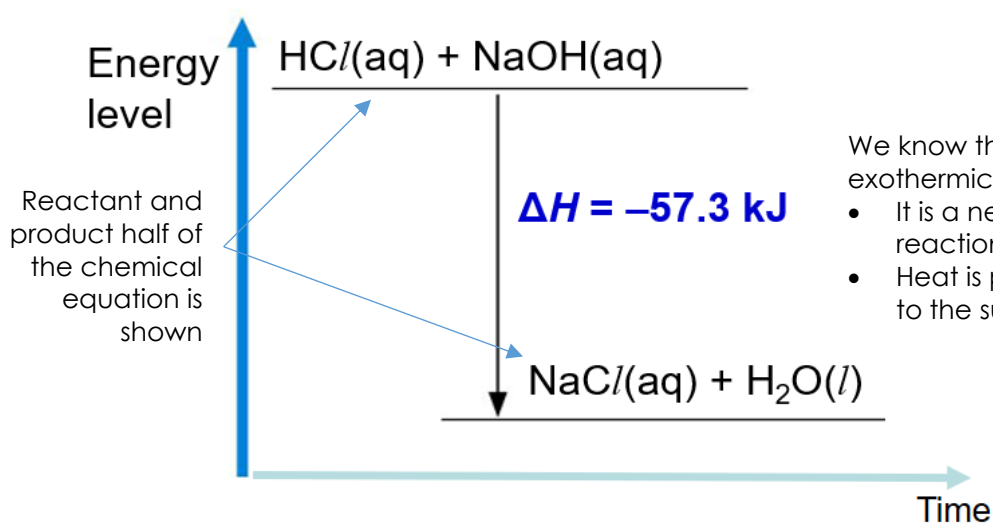




Exothermic Reaction	Endothermic Reaction
Heat is given out to the surroundings.	Heat is taken in from the surroundings.
This means that the total energy of the products is less than that of the reactants.	This means that the total energy of the products is more than that of the reactants.
$\Delta H = \text{Negative}$	$\Delta H = \text{Positive}$

Example

When 1 mol of hydrochloric acid reacts with 1 mol of sodium hydroxide, 57.3 kJ of heat is produced. Draw an energy level diagram for this reaction.



We know that the reaction is exothermic because:

- It is a neutralization reaction
- Heat is produced/released to the surroundings



Application for Calculations

Example 1:



How much heat is given out in

- (i) the combustion of 1 mole of magnesium atoms?
- (ii) burning magnesium to form 20g of MgO?

From the equation, it can be inferred that the reaction involving 2 mol of Mg, 1 mol of O₂ to produce 2 mol of MgO gives out 1200 kJ.

- (i) $1200 \text{ kJ} / 2 = \underline{600 \text{ kJ}}$
- (ii) no. of moles of MgO: $\text{mass}/M_r = 20/(24+16) = 0.5 \text{ mol}$
2 mol MgO gives out 1200 kJ
0.5 mol MgO gives out $1200/4 = \underline{300 \text{ kJ}}$

Example 2:



Calculate the energy released when 14.2g of chlorine reacts completely with hydrogen

From the equation, it can be inferred that the reaction involving 1 mol of H₂, 1 mol of Cl₂ to produce 2 mol of HCl gives out 184 kJ.

- Number of moles of Cl₂: $\text{mass}/M_r = 14.2 / (35.5 \times 2) = 0.2 \text{ mol}$
- 1 mol of Cl₂ gives out 184 kJ
- 0.2 mol of Cl₂ gives out $184/5 = \underline{36.8 \text{ kJ of energy}}$

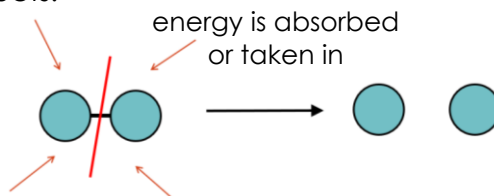


What Causes Energy Changes in Chemical Reactions?

- Reactions involve the **breaking of bonds** and **forming of new bonds**. Thus, **energy changes are caused by the breaking and making of chemical bonds**.

Bond Breaking

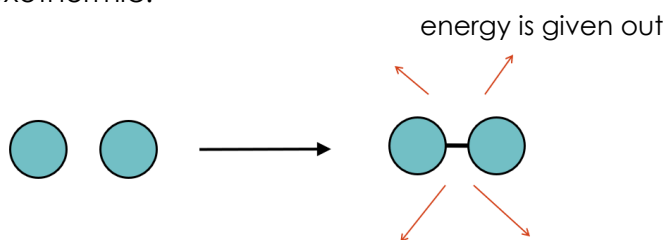
- Bonds in reactant molecules must be broken first before new bonds can be formed in the products.



- Energy is required to break bonds.
- Bond breaking is endothermic

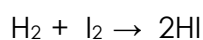
Bond Making/Forming

- Energy is released when bonds are formed.
- Bond making is exothermic.



Example

Bond breaking (in reactants) and bond making (to form products) occurs during chemical reactions.



- H-H and I-I bonds are broken.
- HI bonds are formed.

A chemical reaction is exothermic when:	A chemical reaction is endothermic when:
energy absorbed for bond breaking is less than energy released in bond making	energy absorbed for bond breaking is more than energy released in bond making
$\Delta H_{\text{bond breaking}} < \Delta H_{\text{bond making}}$	$\Delta H_{\text{bond breaking}} > \Delta H_{\text{bond making}}$



Bond Energy – Energy needed to break a chemical bond or energy given out to form the same bond.

Average Bond Enthalpies (kJ/mol)

Single Bonds

C—H 413	N—H 391	O—H 463	F—F 155
C—C 348	N—N 163	O—O 146	
C—N 293	N—O 201	O—F 190	Cl—F 253
C—O 358	N—F 272	O—Cl 203	Cl—Cl 242
C—F 485	N—Cl 200	O—I 234	
C—Cl 328	N—Br 243		Br—F 237
C—Br 276		S—H 339	Br—Cl 218
C—I 240	H—H 436	S—F 327	Br—Br 193
C—S 259	H—F 567	S—Cl 253	
	H—Cl 431	S—Br 218	I—Cl 208
Si—H 323	H—Br 366	S—S 266	I—Br 175
Si—Si 226	H—I 299		I—I 151
Si—C 301			
Si—O 368			

Multiple Bonds

C=C 614	N=N 418	O ₂ 495
C≡C 839	N≡N 941	
C=N 615		S=O 523
C≡N 891		S=S 418
C=O 799		
C≡O 1072		

Calculation using Bond Energy

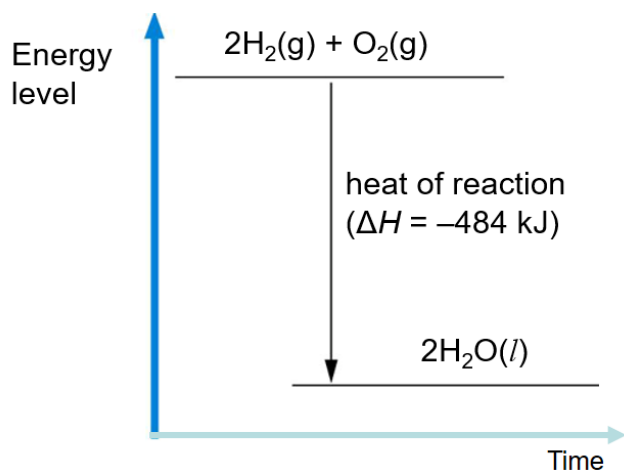
Is the reaction $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$ exothermic or endothermic?

Bond Breaking	Bond Making
Energy absorbed to break 1 mol of H—H bonds = 436 kJ (<i>Bond energy of H-H bond</i>)	Each molecule of H_2O is made up of 2 O—H bonds.
Energy absorbed to break 2 mol of H—H bonds = $2 \times 436 \text{ kJ} = 872 \text{ kJ}$	Energy released on forming 1 mol of O—H bonds = 463 kJ (<i>Bond energy of O-H bond</i>)
Energy absorbed to break 1 mol of O=O bonds = 496 kJ (<i>Bond energy of O=O bond</i>)	Energy released on forming 4 mol of O—H bonds = $4 \times 463 \text{ kJ} = 1852 \text{ kJ}$
Total energy absorbed for bond breaking = $872 + 496 = \underline{\underline{1368 \text{ kJ}}}$	Total energy released for bond making = <u>1852 kJ</u>

$\Delta H_{\text{reaction}} = \text{Heat involved in bond breaking} + \text{Heat involved in bond making}$

$$= 1368 + (-1852) = \underline{\underline{-484 \text{ kJ}}} \quad (< 0)$$

- This reaction is **exothermic**. More energy is **given out to form the O-H bonds** in water (product) than taken in to **break the H-H and O=O** bonds in hydrogen and oxygen gas (reactants).



The energy level diagram tells us that

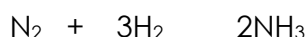
The reaction is exothermic.

The product has less energy than the reactants.

The energy required to break the bonds of H_2 and O_2 molecules is less than the energy released when H_2O bonds are formed.

Example 1:

Is the following reaction exothermic or endothermic? Refer Pg 53



Bond Breaking	Bond Making
Energy absorbed to break 1 mol of $\text{N}\equiv\text{N}$ bonds = 941 kJ (<i>Bond energy of $\text{N}\equiv\text{N}$ bond</i>)	Each molecule of NH_3 is made up of 3 N-H bonds.
Energy absorbed to break 1 mol of H-H bond = 436 kJ (<i>Bond energy of H-H bond</i>)	Energy released on forming 1 mol of N-H bonds = 391 kJ (<i>Bond energy of N-H bond</i>)
Energy absorbed to break 3 mol of H-H bonds = $3 \times 436 \text{ kJ} = 1308 \text{ kJ}$	Energy released on forming 6 mol of N-H bonds = $6 \times 391 \text{ kJ} = 2346 \text{ kJ}$
Total energy absorbed for bond breaking = $941 + 1308 = \underline{\underline{2249 \text{ kJ}}}$	Total energy released for bond making = <u>2346 kJ</u>

$\Delta H_{\text{reaction}} = \text{Heat involved in bond breaking} + \text{Heat involved in bond making}$

$$= 2249 - 2346 = -97 \text{ kJ} \quad (< 0)$$

Hence, the reaction is exothermic. More energy is given out to form the N-H bonds in ammonia (product) than taken in to break the $\text{N}\equiv\text{N}$ and H-H bonds in nitrogen and hydrogen gas (reactants).



Activation Energy

All reactions need energy in order to get started.

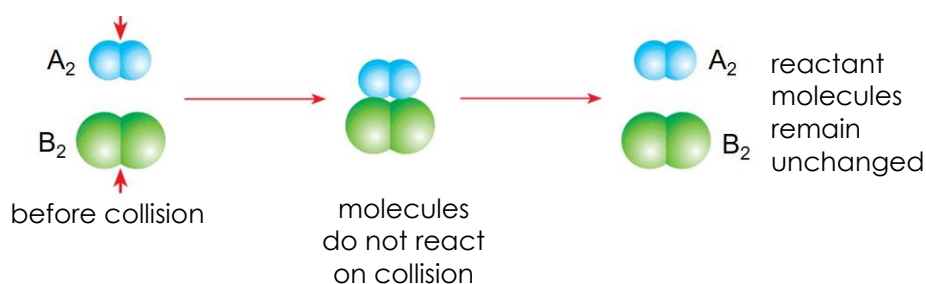
Many reactions that do not occur at room temperature will proceed rapidly when the reactants are heated.

Heat provides the energy to change less energetic particles into more energetic particles and thus start a reaction

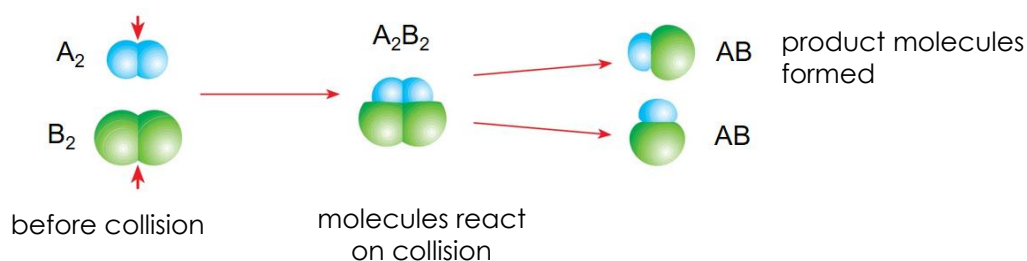
Activation energy, E_a , is the minimum amount of energy that reactant particles must possess in order for a chemical reaction to occur.

A reaction will only occur when reactant particles possess energy equal to or more than the activation energy E_a .

When does a reaction occur?



- Reactant particles have **energy less than E_a** .
- **No reaction** occurs.



- Reactant particles have **energy equal to or greater than E_a** .
- A chemical **reaction** occurs.

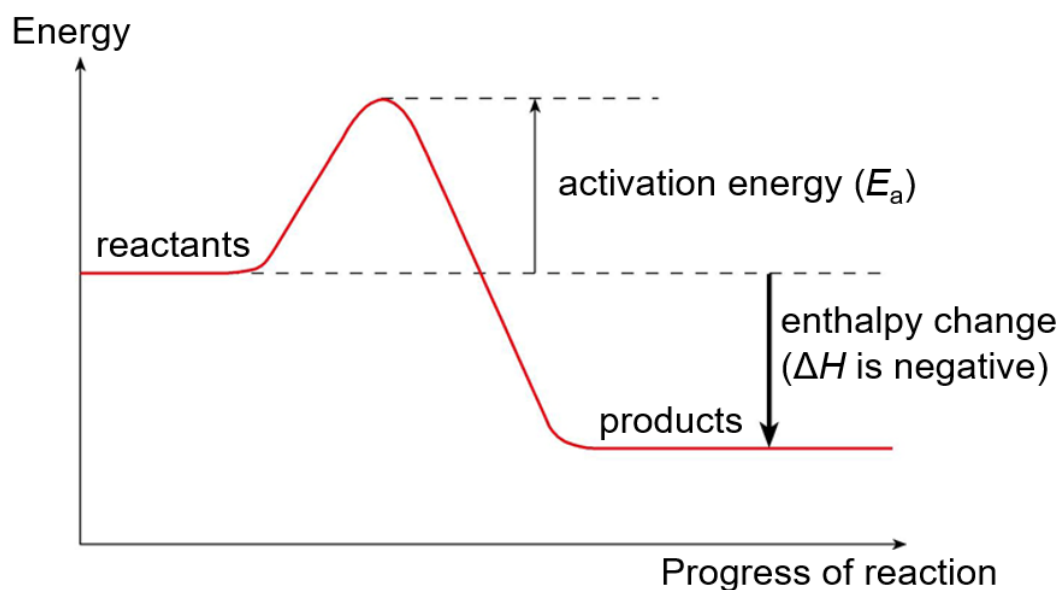


Energy Profile Diagrams

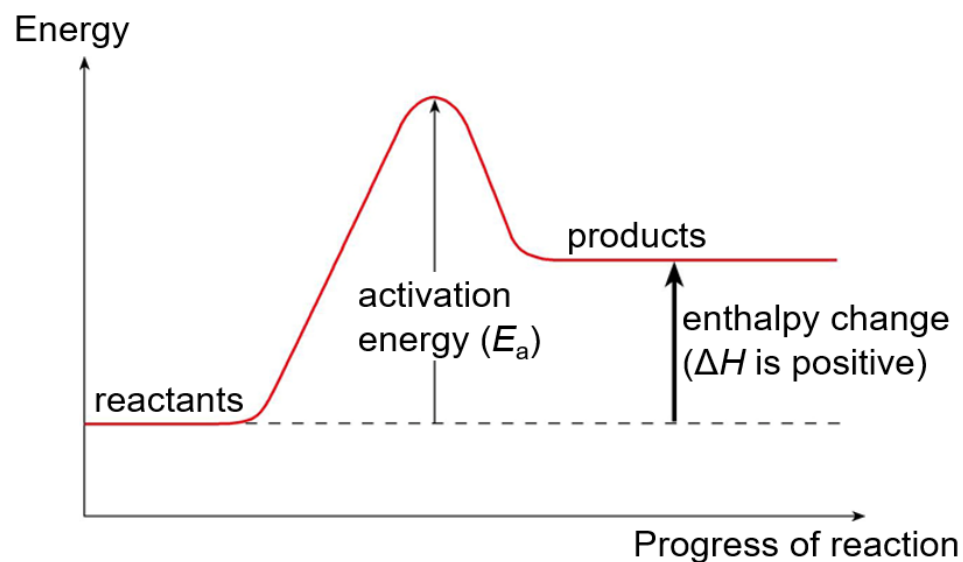
Energy profile diagrams are used to show the **enthalpy change** and **activation energy** of a reaction.

Activation energy acts as an **energy barrier** that must be overcome by the reactants before they can react to form products.

Energy Profile Diagram for An Exothermic Reaction



Energy Profile Diagram for An Endothermic Reaction

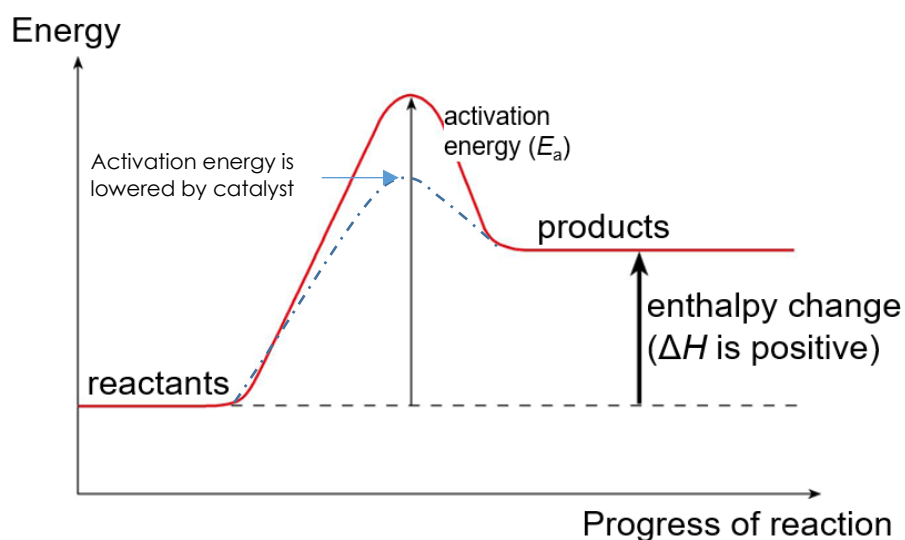
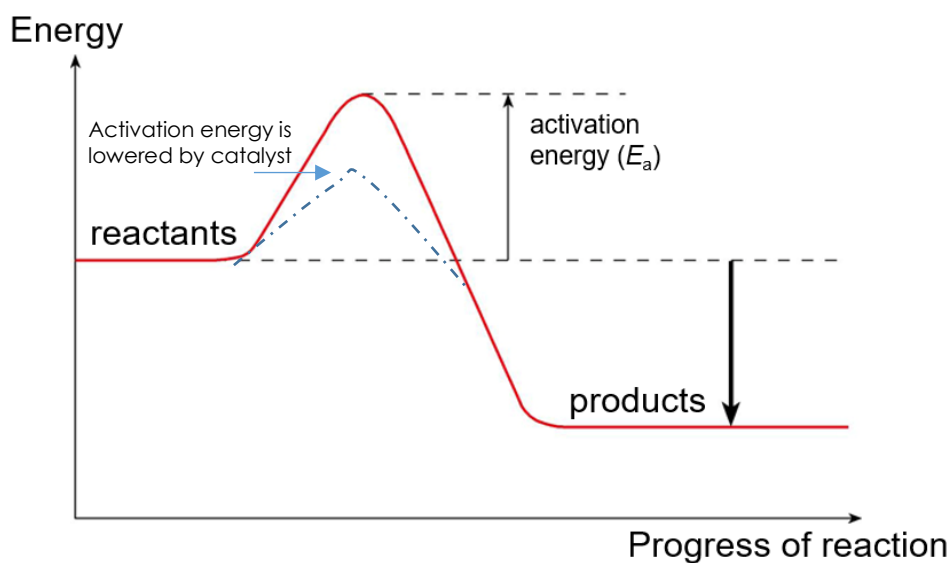




Effect of Catalyst on activation energy

A catalyst

- **speeds up** the rate of a chemical reaction
- remains **chemically unchanged** at the end of the reaction
- provides an **alternative** pathway with a **lower activation energy** for the reaction to take place
- **More reactant particles** now have energy **equal to or greater** than the **activation energy**





Electrolysis

Definition of Electrolysis

The decomposition of a compound by the passage of electricity through it.

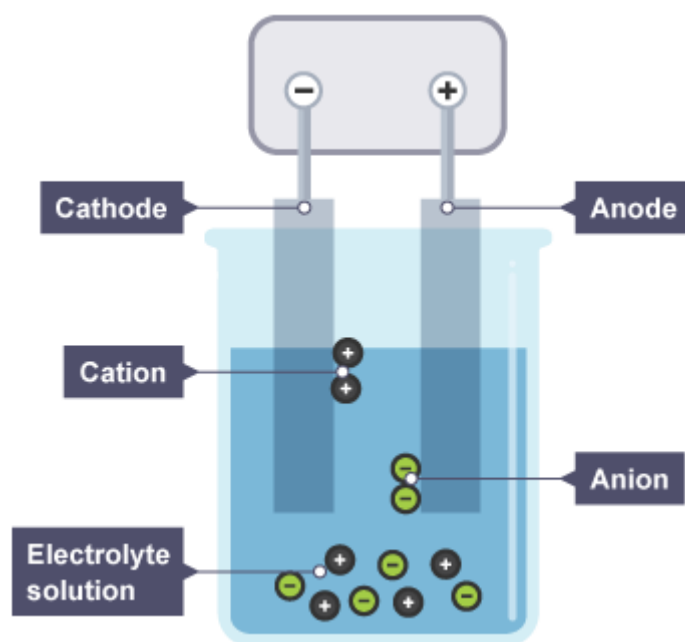
Definition of:

Electrolyte: A **chemical** compound that conducts electricity by changing into ions when melted or dissolved into a solution.

Electrode: A solid electric conductor that carries electric current.

- Cathode: A type of electrode that delivers electrons to cations (positive ions)
- Anode: A type of electrode that receives electrons from anions (negative ions)

PROCESS OF ELECTROLYSIS



Electrons transfer from the negative terminal of the battery and enter the cathode.

At the cathode,

The positive cations (+) are discharged by gaining electrons.

The positive cations become neutral (meta) atoms



At the anode

The anions discharged (-) are discharged by losing electrons.

They become neutral atoms of non-metals.

TYPES OF ELECTROLYTES

- Molten electrolyte
- Aqueous electrolyte

FOR MOLTEN ELECTROLYTES, (Inactive electrode)

Example 1	Molten Lead (II) bromide	Example 2	Molten Sodium chloride
Electrolyte	PbBr_2 (l)	Electrolyte	NaCl (l)
Ions present	Pb^{2+} , Br^-	Ions present	Na^+ , Cl^-
Cathode	$\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$ (l) Pb^{2+} cations are discharged by gaining electrons from the negative terminal of the battery. Neutral Pb metal is formed.	Cathode	$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$ (l) Na^+ cations are discharged by gaining electrons from the negative terminal of the battery. Neutral Na metal is formed.
Anode	$2\text{Br}^- - 2\text{e}^- \rightarrow \text{Br}_2$ (g) Br^- anions are discharged with the loss of electrons to the positive terminal of the battery. Br atoms are formed. Each Br atom covalently bonds to another Br atom to form Br_2 (g)		$2\text{Cl}^- - 2\text{e}^- \rightarrow \text{Cl}_2$ (g) Cl^- anions are discharged with the loss of electrons to the positive terminal of the battery. Cl atoms are formed. Each Cl atom covalently bonds to another Cl atom to form Cl_2 (g)
Observation	Lead is found as a molten globules (liquid) at the bottom of the crucible. Reddish brown bromine gas is formed at the anode.	Observation	Sodium is found as a molten globules (liquid) at the bottom of the crucible. Yellow chlorine gas is formed at the anode.
Overall reaction	PbBr_2 (l) \rightarrow Pb (l) + Br_2 (g)	Overall reaction	2NaCl (l) \rightarrow 2Na (l) + Cl_2 (g)



FOR AQUEOUS ELECTROLYTES,

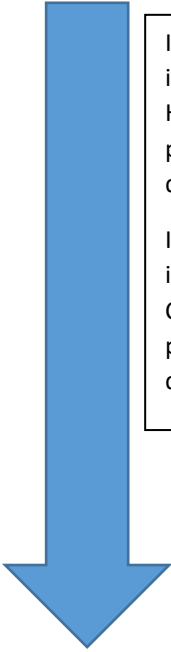
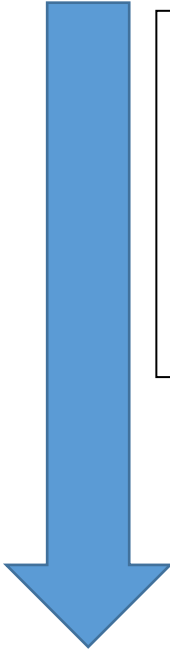
- Dilute aqueous solution
- Concentrated aqueous solution

For both cases, water is present. The ions of water, H^+ and OH^- , are to be considered during electrolysis. The rule for preferential discharge comes into play.

RULE FOR PREFERENTIAL DISCHARGE

It states that if more than one type of **ions** are attracted towards a particular electrode, then the one **discharged** is the **ion** which requires least energy.

- The cation (less reactive metal ion) has a lowest position in the electrochemical series are preferentially **discharged**.
- The anion has a lowest position in the electrochemical series are preferentially **discharged**.

Ions	Ease of discharge	Ions	Ease of discharge
K^+	Harder to discharge	SO_4^{2-}	Harder to discharge
		NO_3^-	
Na^+ Ca^{2+} Mg^{2+} Al^{3+} Zn^{2+} Fe^{2+} Pb^{2+} H^+ Cu^{2+} Hg^{2+}	 <div> <p>If both H^+ and Na^+ ions are present, H^+ will be preferentially discharged.</p> <p>If both H^+ and Cu^{2+} ions are present, Cu^{2+} will be preferentially discharged.</p> </div>	F^- Cl^- Br^- I^-	 <div> <p>If both NO_3^- and OH^- are present, OH^- will be preferentially discharged.</p> <p>(Note: Sulfates and Nitrates are not discharged.)</p> </div>
Ag^+	Easiest to discharge	OH^-	Easiest to discharge

For both NO_3^- and SO_4^{2-} , they are not discharged. The Nitrogen atom in Nitrate ions has an oxidation state of +5. To discharge the nitrate ion, you have to oxidise it, but it's very difficult to oxidise the Nitrogen atom from +5 to a +6 as Nitrogen has only 5 valence electrons. You got to take an electron from an inner shell which requires lots of energy.

This is similar for sulfate ions, where you can't oxidise the S atom from a +6 to a +7 or +8.


FOR DILUTED AQUEOUS ELECTROLYTES, (Inactive electrode)

Example 1	Dilute sulfuric acid	Example 2	Aqueous copper(II) chloride
Electrolyte	H ₂ SO ₄ (aq)	Electrolyte	CuCl ₂ (aq)
Ions present	H ⁺ , SO ₄ ²⁻ , OH ⁻ Water contains H ⁺ & OH ⁻ ions	Ions present	Cu ²⁺ , Cl ⁻ H ⁺ , OH ⁻ Water contains H ⁺ & OH ⁻ ions
Cathode(Inert)	$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2 (\text{g})$ H ⁺ cations are discharged by gaining electrons from the negative terminal of the battery. Hydrogen gas is formed.	Cathode(Inert)	<ul style="list-style-type: none"> Cu²⁺ is preferentially discharged H⁺ is not discharged $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu} (\text{l})$ Cu ²⁺ cations are discharged by gaining electrons from the negative terminal of the battery. Neutral Cu metal is formed.
Anode(Inert)	<ul style="list-style-type: none"> OH⁻ is preferentially discharged SO₄²⁻ is not discharged $4\text{OH}^- - 4\text{e}^- \rightarrow \text{O}_2 (\text{g}) + 2\text{H}_2\text{O} (\text{l})$ OH ⁻ anions are discharged with the loss of electrons to the positive terminal of the battery. O ₂ and H ₂ O molecules are formed.	Anode(Inert)	<ul style="list-style-type: none"> OH⁻ is preferentially discharged Cl⁻ is not discharged $4\text{OH}^- - 4\text{e}^- \rightarrow \text{O}_2 (\text{g}) + 2\text{H}_2\text{O} (\text{l})$ OH ⁻ anions are discharged with the loss of electrons to the positive terminal of the battery. O ₂ and H ₂ O molecules are formed.
Observation	Bubbling occurs at both the cathode and anode. Hydrogen is produced at the cathode and oxygen at the anode. pH of the electrolyte remains the same. $4\text{H}^+ (\text{aq}) + 4\text{OH}^- (\text{aq}) \rightarrow 2\text{H}_2 (\text{g}) + \text{O}_2 (\text{g}) + 2\text{H}_2\text{O} (\text{l})$ Note: The volume of hydrogen gas is doubled that of oxygen.	Observation	Reddish brown solid copper is deposited at the cathode. Bubbling occurs at the anode, oxygen gas is produced. pH of electrolyte decreases as more OH ⁻ are discharged leaving H ⁺ ions not discharge. Since there is lesser Cu ²⁺ in solution, blue coloured CuSO ₄ fades $2\text{Cu}^{2+} (\text{aq}) + 4\text{OH}^- (\text{aq}) \rightarrow 2\text{Cu} (\text{s}) + \text{O}_2 (\text{g}) + 2\text{H}_2\text{O} (\text{l})$



DIAGRAMS

Diagram showing process of electrolysis of molten lead (II) bromide

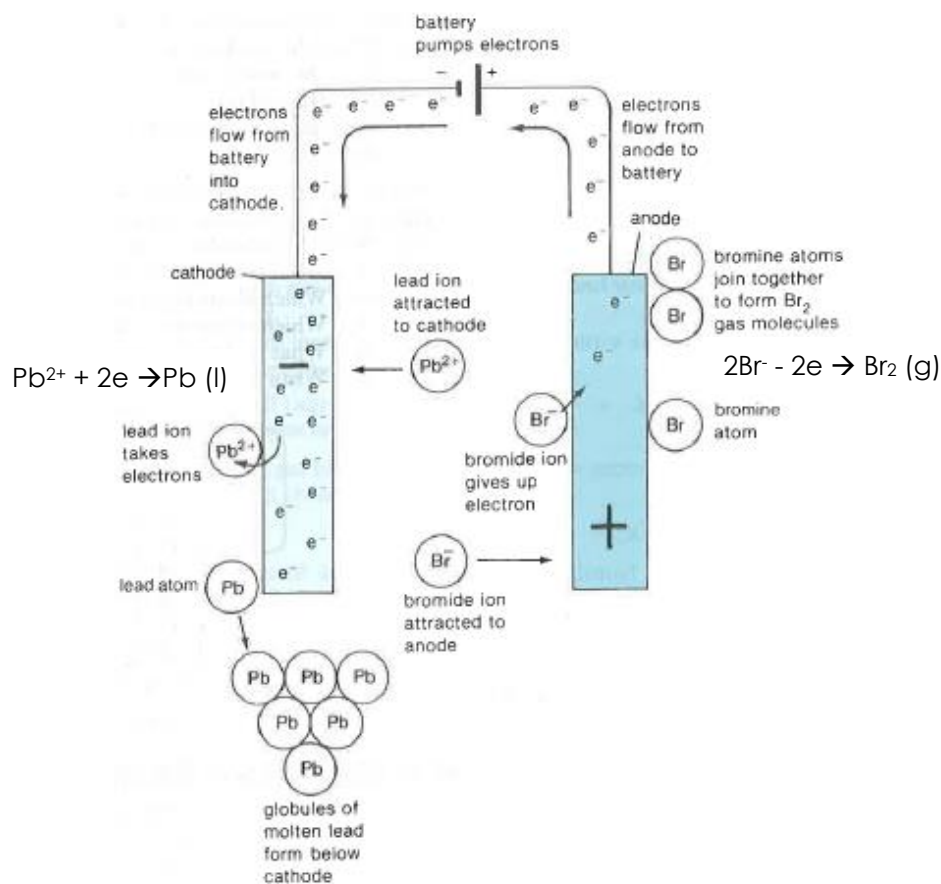
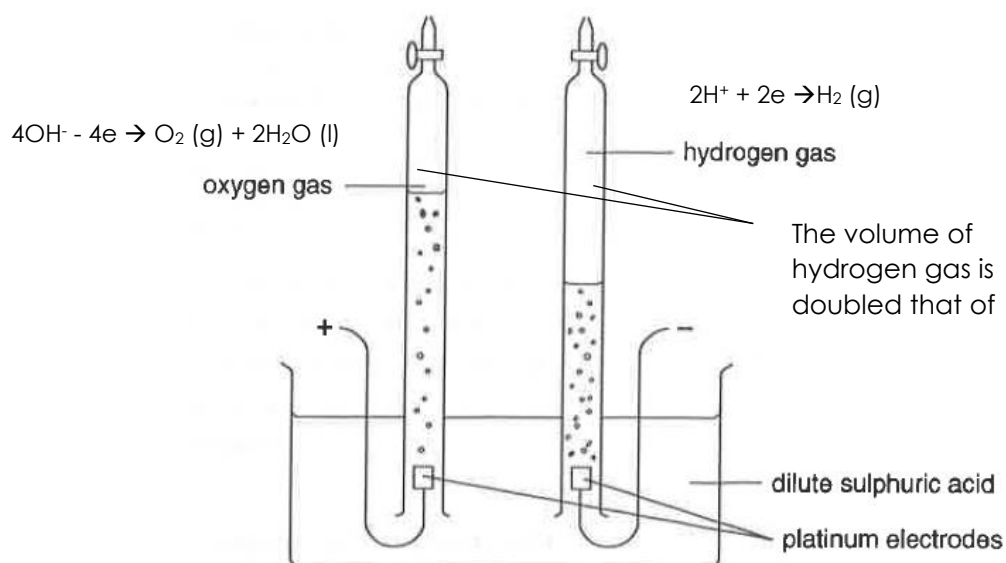


Diagram showing process of electrolysis of dilute sulfuric acid




FOR CONCENTRATED AQUEOUS ELECTROLYTES, (Inactive electrode)

Example 1	Concentrated Sodium chloride solution (Brine)	Example 2	Concentrated Copper(II) chloride solution
Electrolyte	Conc NaCl (aq)	Electrolyte	Conc CuCl ₂ (aq)
Ions present	H ⁺ , OH ⁻ (Higher Conc) : Na ⁺ , Cl ⁻	Ions present	H ⁺ , OH ⁻ (Higher Conc) : Cu ²⁺ , Cl ⁻
Cathode(Inert)	Na ⁺ and H ⁺ are attracted As Na ⁺ is more stable, H ⁺ cations are discharged. $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2 (\text{g})$ Hydrogen gas is formed.	Cathode(Inert)	Cu ²⁺ and H ⁺ are attracted <ul style="list-style-type: none"> Cu²⁺ is preferentially discharged H⁺ is not discharged $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu} (\text{l})$ Cu ²⁺ cations are discharged by gaining electrons from the negative terminal of the battery. Neutral Cu metal is formed.
Anode (Inert)	OH ⁻ and Cl ⁻ are attracted. <ul style="list-style-type: none"> OH⁻ is should be discharged but Cl⁻ is present in a greater quantity. Cl⁻ is discharged as Cl₂ $2\text{Cl}^- - 2\text{e}^- \rightarrow \text{Cl}_2 (\text{g})$ Cl ⁻ anions are discharged. Cl ₂ is formed.	Anode(Inert)	OH ⁻ and Cl ⁻ are attracted. <ul style="list-style-type: none"> OH⁻ is should be discharged but Cl⁻ is present in a greater quantity. Cl⁻ is discharged as Cl₂ $2\text{Cl}^- - 2\text{e}^- \rightarrow \text{Cl}_2 (\text{g})$ Cl ⁻ anions are discharged. Cl ₂ is formed.
Observation	Bubbling occurs at both the cathode and anode. Hydrogen is produced at the cathode and chlorine gas at the anode. As H ⁺ is removed as H ₂ gas, the concentration of undischarged OH ⁻ increases relatively. pH of the electrolyte increases. Solution becomes more alkaline.	Observation	Reddish brown solid copper is deposited at the cathode. Bubbling occurs at the anode, chlorine gas is produced. Since there is lesser Cu ²⁺ in solution, blue coloured CuCl ₂ fades pH remain constant as no H ⁺ or OH ⁻ ions are discharged.



ACTIVE ELECTRODES

Inert electrodes such as carbon and platinum do not dissolve into the electrolyte.

When active electrodes such as copper or silver are used, they may dissolve into the electrolyte and result in changes in the products formed.

Compare the reactions for when an active and inert electrodes are used.

Example 1a	Aqueous Silver nitrate solution	Example 1b	Aqueous Silver nitrate solution
Electrolyte	AgNO_3 (aq)	Electrolyte	AgNO_3 (aq)
Ions present	H^+ , OH^- Ag^+ , NO_3^-	Ions present	H^+ , OH^- Ag^+ , NO_3^-
Cathode (Inert)	Ag^+ and H^+ are attracted Ag^+ is more preferentially discharged. $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag (s)}$ Silver is deposited on the cathode	Cathode (Silver cathode)	Ag^+ and H^+ are attracted Ag^+ is more preferentially discharged. $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag (s)}$ Silver is deposited on the cathode
Anode (Inert)	OH^- and NO_3^- are attracted. <ul style="list-style-type: none"> OH^- is discharged. $4\text{OH}^- - 4\text{e}^- \rightarrow \text{O}_2 (\text{g}) + 2\text{H}_2\text{O} (\text{l})$ OH^- anions are discharged. O_2 and H_2O molecules are formed.	Anode (Silver anode)	OH^- and NO_3^- are attracted. But Ag anode dissolves into the electrolyte to form Ag^+ (aq). $\text{Ag} - \text{e}^- \rightarrow \text{Ag}^+ (\text{aq})$ Ag^+ then migrate towards the cathode. The mass of the anode decreases. The mass of the cathode increases as Ag^+ is discharged there.
Observation	Bubbling occurs at both the cathode and anode. Hydrogen is produced at the cathode and oxygen gas at the anode. As OH^- is removed, the concentration of H^+ increases relatively. pH of the electrolyte decreases. Solution becomes more acidic.	Observation	Cathode increases in mass Anode dissolves into the electrolyte and its mass decreases. No gas is produced.

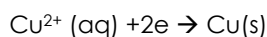


Example 2a	Aqueous Copper (II) sulfate solution	Example 2b	Aqueous Copper (II) sulfate solution
Electrolyte	CuSO_4 (aq)	Electrolyte	CuSO_4 (aq)
Ions present	H^+ , OH^- Cu^{2+} , SO_4^{2-}	Ions present	H^+ , OH^- Cu^{2+} , SO_4^{2-}
Cathode (Inert)	<p>Cu^{2+} and H^+ are attracted</p> <p>Cu^{2+} is more preferentially discharged.</p> <p>$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu (s)}$</p> <p>Copper is deposited on the cathode</p>	Cathode (copper cathode)	<p>Cu^{2+} and H^+ are attracted</p> <p>Cu^{2+} is more preferentially discharged.</p> <p>$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu (s)}$</p> <p>Copper is deposited on the cathode</p>
Anode (Inert)	<p>OH^- and SO_4^{2-} are attracted.</p> <ul style="list-style-type: none"> OH^- is discharged. <p>$4\text{OH}^- - 4\text{e}^- \rightarrow \text{O}_2 \text{ (g)} + 2\text{H}_2\text{O (l)}$</p> <p>$\text{OH}^-$ anions are discharged. O_2 and H_2O molecules are formed.</p>	Anode (copper anode)	<p>OH^- and SO_4^{2-} are attracted.</p> <p>But Cu anode dissolves into the electrolyte to form Cu^{2+} (aq).</p> <p>$\text{Cu (s)} \rightarrow \text{Cu}^{2+} \text{ (aq)} + 2\text{e}^-$</p> <p>$\text{Cu}^{2+}$ then migrate towards the cathode. The mass of the anode decreases. The mass of the cathode increases as Cu^{2+} is discharged there.</p>
Observation	<p>Bubbling occurs at the anode. Oxygen gas at the anode.</p> <p>As OH^- is removed, the concentration of H^+ increases relatively. pH of the electrolyte decreases. Solution becomes more acidic.</p> <p>Copper is discharged at the cathode. The blue solution fades in colour.</p>	Observation	<p>Cathode increases in mass Anode dissolves into the electrolyte and its mass decreases.</p> <p>No gas is produced.</p> <p>The colour of the solution does not fade. For each Cu^{2+} discharged at the cathode, a Cu^{2+} is released into the electrolyte from the anode.</p>

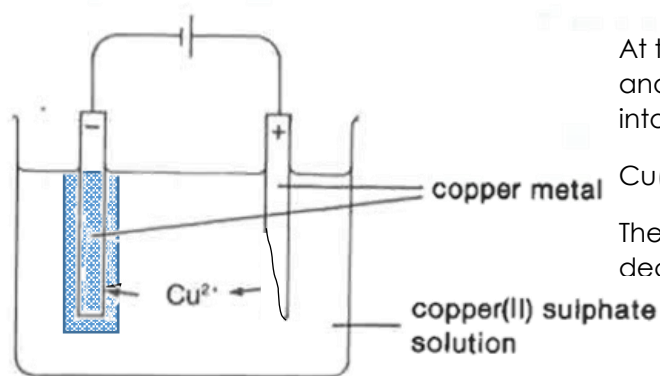


DIAGRAMS

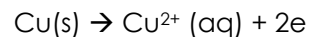
At the cathode, Cu^{2+} ions migrate towards it and discharged to form Cu metal.



The mass of cathode increases.



At the Anode, Cu anode itself dissolves into the electrolyte.

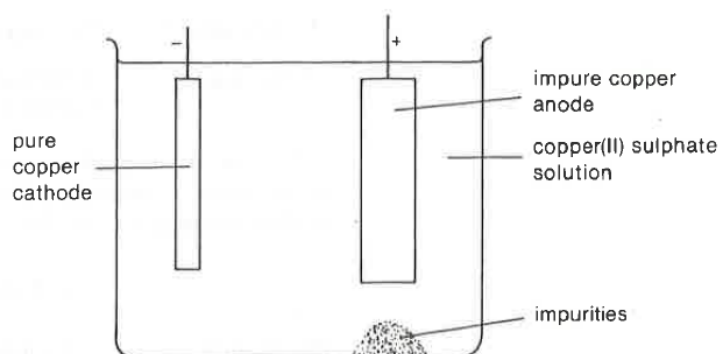


The mass of the anode decreases.

The colour of the solution remains the same as Cu^{2+} ions is released into the electrolyte from the anode.

APPLICATIONS USING ACTIVE ELECTRODES

- Purification of Metal**

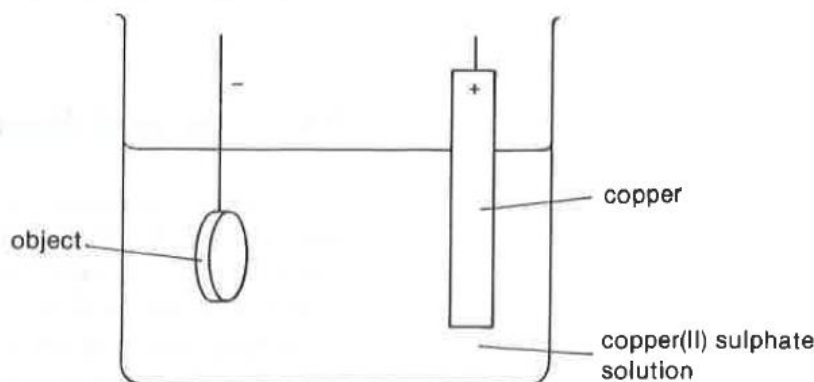


To purify a metal eg. Copper, the impure copper is made the anode. A piece of pure copper is used as the cathode. Pure copper from the anode dissolves into the electrolyte. Pure copper is then deposited at the cathode. The impurities left on the anode, fall off and can be collected below the anode.



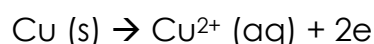
- **Electro-plating**

A metallic object can be coated with another metal by electrolysis. This process is known as electroplating.



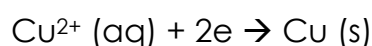
The electrolyte used must have the same positive ions as the metal to be coated. To electroplate an object with copper, copper (II) sulfate solution is used as the electrolyte. The object is made the cathode and pure copper is used as the anode.

At the anode, the copper metal dissolves.



The Cu^{2+} ions migrate towards the cathode

At the cathode, the Cu^{2+} ions are discharged.



Copper metal is deposited on the object. The object eventually gets electroplated or covered with copper metal.

Steel objects are electroplated to protect the steel from corrosion and to give the object an attractive shiny appearance. The metals used to electroplate are usually corrosion resistant, shiny and low in the reactivity series. Below are some examples of electroplating.

Electroplating metal	Uses
Chromium	Water taps, motorcar and bicycle parts
Tin	Tin cans
Silver	Silver-plated sports trophies, plaques, ornaments, knives and forks (silverwares)
Nickel	Corrosion resistant layer before electroplating with another metal on top.



EXTRACTION OF METALS

Potassium (K)	Electrolysis	
Sodium (Na)		
Calcium (Ca)		
Magnesium (Mg)		
Aluminium (Al)		
Zinc (Zn)	Reduction with carbon	
Iron (Fe)		
Lead (Pb)		
Copper (Cu)		
Silver (Ag)		
Gold (Au)	Occurs naturally uncombined.	
	Physical mining (excavation)	

- Metals can be extracted from their ores by electrolysis. However this process is very energy consuming and expensive. This method of extraction is only used to extract very reactive metal such as sodium, calcium and aluminium.
- These metals cannot be extracted from their ores by reduction with carbon.
- The pure metal can only be obtained by electrolysing a molten ionic compound of the metal.

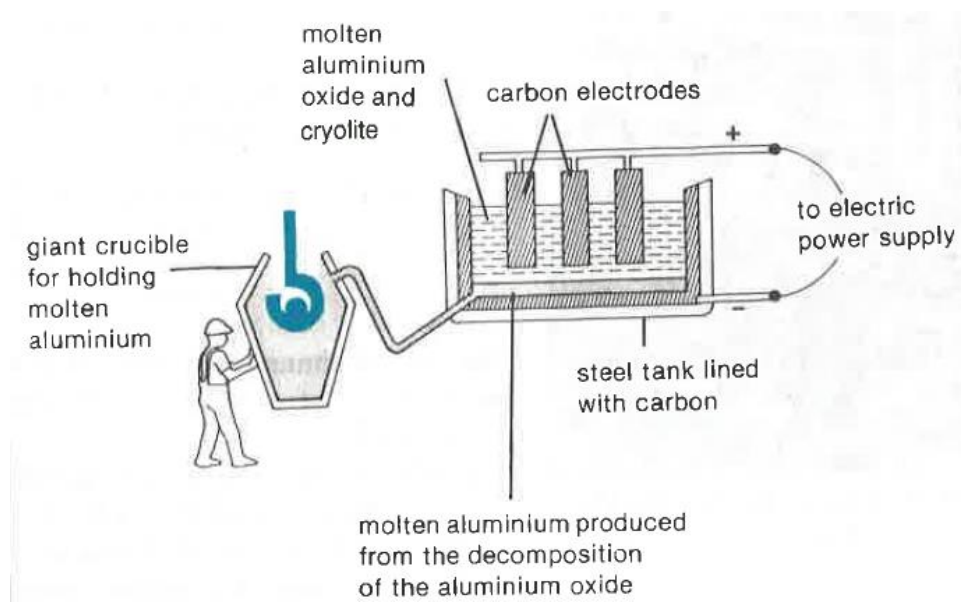


EXTRACTION OF ALUMINIUM

Ore: Bauxite (Al_2O_3)

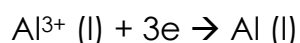
Melting point: 2000°C

Catalyst (Cryolite) is used to lower the mp to 900°C



An electric current is passed through molten mixture of aluminium oxide and cryolite. Inert carbon electrodes are used for both cathode and anode.

At the cathode, the Al^{3+} ions are discharged and pure molten aluminium is deposited at the bottom of the steel tank.



At the anode, the O^{2-} ions are discharged and oxygen gas is produced. However the hot oxygen burns the carbon anodes away and the carbon anodes need to be replaced.



Overall Reaction: $2\text{Al}_2\text{O}_3 (\text{l}) \rightarrow 4\text{Al} (\text{l}) + 3\text{O}_2 (\text{g})$



SIMPLE CELLS

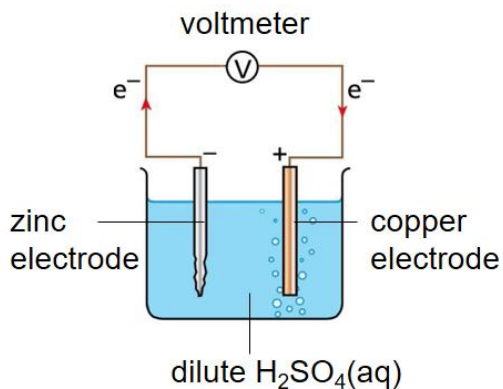
A simple cell is a device that converts chemical energy into electrical energy. It is also known as an electric cell.

A simple cell is made by **placing two different metals in contact with an electrolyte**. The **metals** act as **electrodes** for the simple cell.

ELECTROLYTIC CELL VS SIMPLE CELL

Electrolytic cell	Simple cell
Electrical energy (from battery) is converted into chemical energy	Chemical energy is converted into electrical energy

SIMPLE CELL: HOW IT WORKS?



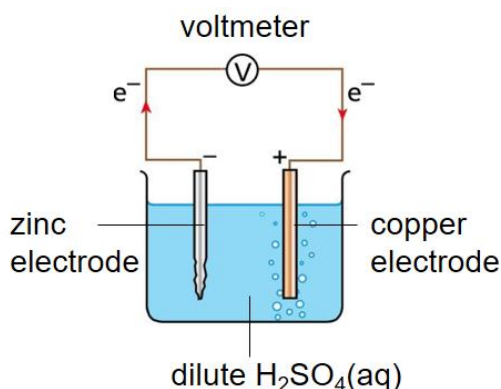
When a zinc plate and a copper plate are placed in dilute sulfuric acid and connected by wires, a potential difference is set up between the metal plates.

The potential difference can be registered by a voltmeter.

Site of REDUCTION: CATHODE	RC
Site of OXIDATION: ANODE	OA



Example 1: Electrolyte is $\text{H}_2\text{SO}_4(\text{aq})$



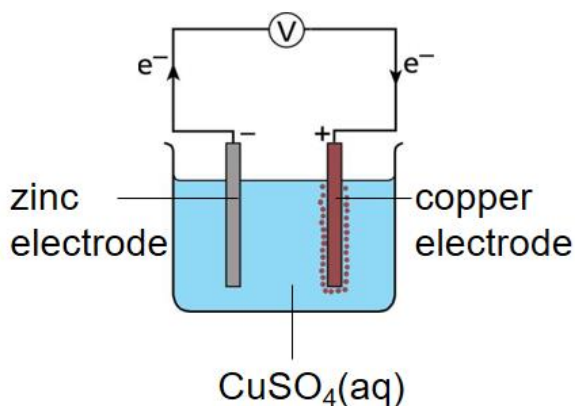
At the zinc electrode:	At the copper electrode:
<p>Zinc atoms (being more reactive than copper) give up electrons and go into solution as zinc ions.</p> $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$ <p>(oxidation)</p> <p>The electrode from which electrons flow out is the negative electrode.</p>	<p>Hydrogen ions take up electrons (released by Zinc) to form hydrogen gas.</p> $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$ <p>(reduction)</p>
<p>Zinc is the negative electrode. (ANODE - OXIDATION SITE)</p> <p>The electrons flow from the zinc electrode along the wire to the copper electrode.</p>	<p>Copper is the positive electrode. (CATHODE - REDUCTION SITE)</p>
<p>The zinc electrode becomes smaller.</p>	<p>Hydrogen gas is formed at the copper electrode</p>

Overall cell reaction: $\text{Zn(s)} + 2\text{H}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_2(\text{g})$

- Oxidation and reduction occur to cause the flow of electrons.
- Hence, electrical energy is produced by redox reactions in a simple cell.
- In a simple cell, the flow of electrons is always from the more reactive metal to the less reactive metal.
- The **more reactive metal** becomes the **negative electrode** and the less reactive metal the positive electrode.



Example 2: Electrolyte is $\text{CuSO}_4(\text{aq})$



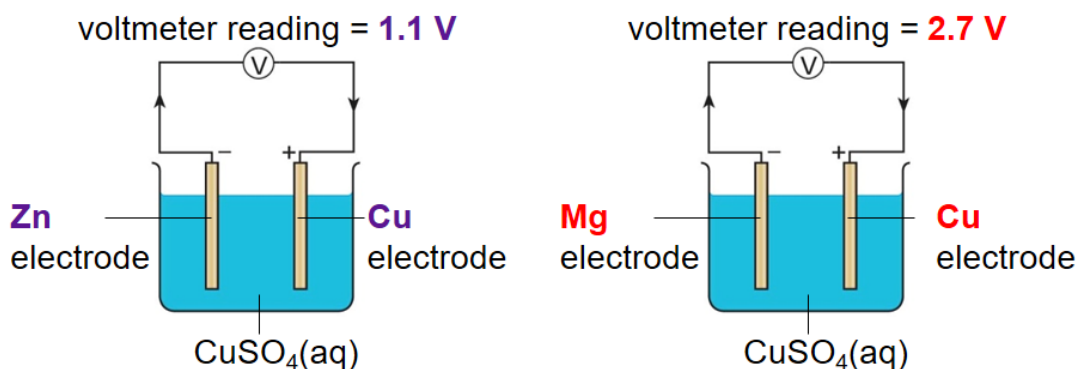
At the zinc electrode:	At the copper electrode:
<p>Zinc atoms (being more reactive than copper) give up electrons and go into solution as zinc ions.</p> $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-}$ <p>(oxidation)</p> <p>The electrode from which electrons flow out is the negative electrode.</p>	<p>Copper(II) ions from the solution receive electrons (released by Zinc) at the copper electrode</p> $2\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Cu(s)}$ <p>(reduction)</p>
<p>Zinc is the negative electrode.</p> <p>(ANODE - OXIDATION SITE)</p> <p>The electrons flow from the zinc electrode along the wire to the copper electrode.</p>	<p>Copper is the positive electrode.</p> <p>(CATHODE - REDUCTION SITE)</p>
<p>The zinc electrode becomes smaller.</p>	<p>Copper electrode gain in mass</p>

Overall ionic equation: $\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$



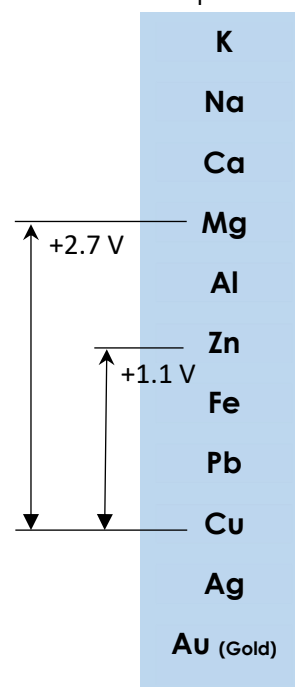
SIMPLE CELL AND THE REACTIVITY SERIES

The amount of **electrical energy** produced in a simple cell is determined by **how far apart the metals** (electrodes) are in the reactivity series.



Different voltages are produced when different metals are used in the simple cell.

Metal electrodes	Voltage/V
magnesium/copper	2.7
zinc/copper	1.1
iron/copper	0.8
lead/copper	0.5
copper/copper	0.0

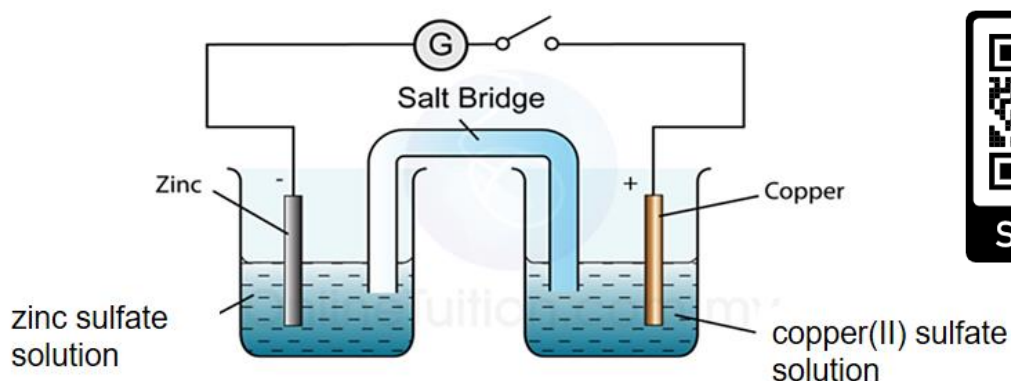


The further apart the two metals are in the reactivity series, the greater the voltage produced.

No current will flow if both electrodes are made of the same metal.



GALVANIC/VOLTAIC CELL AND SALT BRIDGE



- The purpose of the salt bridge is to maintain charge neutrality in each beaker while electrons flow from one electrode to the other.
- To complete the electrical circuit by allowing the ions to flow between the solutions.

Why a salt bridge?

	<p>Zinc sulfate solution Contains Zn^{2+} and SO_4^{2-} ions</p> <p>Copper (II) sulfate solution Contains Cu^{2+} and SO_4^{2-} ions</p>
	<p>Zinc electrode being more electropositive gives away electrons and becomes Zn^{2+} ions in solution</p> <p>Cu^{2+} ions in solution gain electrons to form another layer of copper.</p>



<p>If the build up of charges is too much, the current stops</p> <p>Positive charges build up Net Zn^{2+} ions in solution continues to increase</p> <p>Negative charges build up. More sulfates ions than copper ions Net Cu^{2+} ions in solution decreases.</p>	<p>Net Zn^{2+} ions in solution continues to increase. Positive charges build up.</p> <p>Net Cu^{2+} ions in solution decreases. Negative charges build up as there are relatively more sulfates ions than copper ions now.</p> <p>If the builds up of charges occurs on both sides, electrons transfer will eventually stop and the electrical current will cease.</p>
<p>Salt Bridge Contains NaCl solution</p>	<p>To prevent this, a salt bridge is used to supply ions. A salt bridge like NaCl contains Na^+ and Cl^- ions. A cotton plug is placed at both end to prevent NaCl solution from flowing directly into the half cells.</p>
<p>Salt Bridge Reduces the build up of positive ions and negative ions on both sides. Hence further conduction can occur.</p> <p>Cl^- ions from salt bridge enters into solution thus reducing the positive build up</p> <p>Na^+ ions from salt bridge enters into solution thus reducing the negative build up</p>	<p>Cl^- ions from salt bridge enters into solution thus reducing the positive build up</p> <p>Na^+ ions from salt bridge enters into solution thus reducing the negative build up</p> <p>The Salt Bridge thus reduces the build up of positive ions and negative ions on both sides, enabling further conduction to occur.</p>



SUMMARY

Electrolysis

involves

Passing electricity through an electrolyte (molten or aqueous ionic compound) to decompose it

into

Cations

K^+
 Na^+
 Ca^{2+}
 Mg^{2+}
 Zn^{2+}
 Fe^{2+}
 Pb^{2+}
 H^+
 Cu^{2+}
 Ag^+

Ease of discharge of ion increases

Anions

SO_4^{2-}
 NO_3^-
 F^-
 Cl^-
 Br^-
 I^-
 OH^-

(Note: SO_4^{2-} and NO_3^- are not discharged.)

factors affecting selective discharge

factors affecting selective discharge

Position of metal in the reactivity series depends on

Ease of discharge of cation

Ease of discharge of anion

Concentration of anion in the electrolyte

used for predicting

Electrolyte	Electrodes	Products of electrolysis	
		at anode (+)	at cathode (-)
molten $NaCl(l)$	carbon	$Cl_2(g)$	$Na(l)$
dilute $NaCl(aq)$	carbon	$O_2(g)$	$H_2(g)$
concentrated $NaCl(aq)$	carbon	$Cl_2(g)$	$H_2(g)$
$CuSO_4(aq)$	platinum	$O_2(g)$	$Cu(s)$
$CuSO_4(aq)$	copper	Cu anode dissolves	$Cu(s)$
$CuSO_4(aq)$	Cu (anode), C (cathode)	Cu anode dissolves	$Cu(s)$

Using inert electrodes

Using reactive electrodes

Copper plating
 Anode: pure copper
 Cathode: object to be plated

industrial applications

Purification of copper
 Anode: impure copper
 Cathode: pure copper

Simple Cell

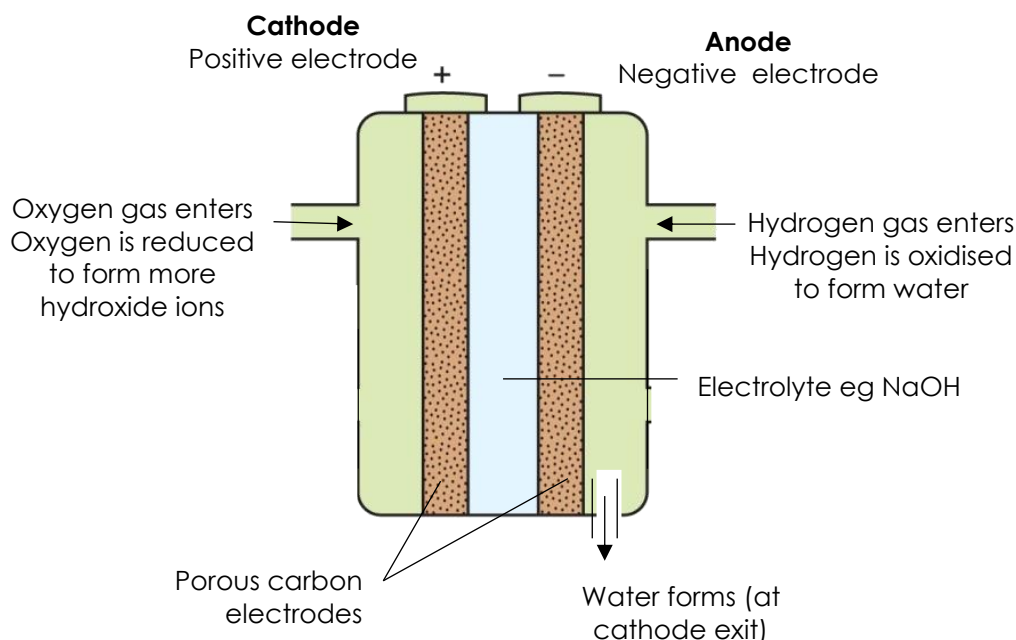
- The more reactive metal becomes the negative electrode; the less reactive metal becomes the positive electrode.
- Flow of electrons from the more reactive metal to the less reactive metal
- The further apart the two metals are in the reactivity series, the higher the voltage produced.



FUEL CELLS

A chemical cell in which the reactants are continuously supplied to produce electrical directly eg. Hydrogen and oxygen fuel cell (used as source of energy in space vehicles)

Parts of a Hydrogen - Oxygen fuel cell



How does a fuel cell work?

At the Anode (-)	<ul style="list-style-type: none"> Hydrogen gas undergoes oxidation by releasing electrons to form H^+ ions. The H^+ ions react with OH^- from the electrolyte to form water. Water formed at the anode is removed. 	$2H_2(g) + 4OH^-(aq) \rightarrow 4H_2O(l) + 4e^-$
At Cathode (+)	<ul style="list-style-type: none"> Oxygen gas undergoes reduction. Oxygen gains electrons and reacts with water to form hydroxide ions. Hydroxide ions formed will react with the hydrogen gas to form water. 	$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
Overall Reaction	$O_2(g) + 2H_2(g) \rightarrow 2H_2O(l)$	$2H_2(g) + 4OH^-(aq) \rightarrow 2H_2O(l) + 4e^-$ $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$ <hr/> $O_2(g) + 2H_2(g) \rightarrow 2H_2O(l)$

**Difference between Fuel and Ordinary Cell**

- The reactants are not contained within the cell but instead are continuously supplied from an external reservoir.

Advantages of using Hydrogen as a Fuel

- It is a renewable energy resource.
- It is also pollution-free because only water is produced when hydrogen reacts with oxygen.
- Hydrogen is an efficient source of energy, which provides twice the amount of energy compared to many other fuels.

Disadvantages of Using Hydrogen As a Fuel

- It is difficult to find a cheap source of hydrogen. Obtaining hydrogen from electrolysis is an expensive process.
- Hydrogen is extremely flammable and explosive. Special precautions must be taken in the storage and transport of the gas.
- Hydrogen is very light, and thus requires larger storage volumes compared to other fuels.





AIR

PART I: Composition of Air

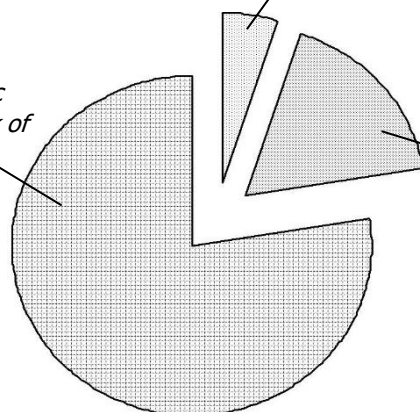
Name of gas	Formula	% composition in air	Type of gas
Nitrogen	N ₂	78	Neutral
Oxygen	O ₂	21	Neutral
Carbon dioxide	CO ₂	0.03	Acidic gas
Rare gases	mainly Ar	0.97	Noble gas

1% Other Gases

Mainly argon, with some carbon dioxide, and traces of other noble gases.

78% Nitrogen (N₂)

A relatively inert diatomic gas which forms the bulk of our atmosphere.



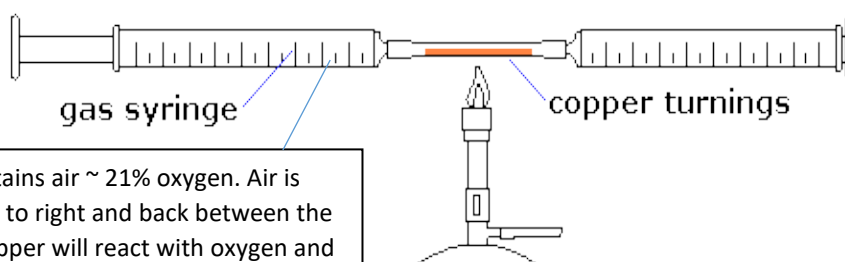
21% Oxygen (O₂)

Plays a vital role in many processes, such as respiration and combustion.

PART II: Questions involving the percentage of oxygen in air.

Example 1: Oxidation of Copper (Syringe Method)

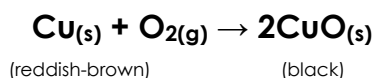
The volume of oxygen present in air can be found by passing a fixed volume of air over hot copper turnings, as shown in the diagram below,



Gas syringe contains air ~ 21% oxygen. Air is passed from left to right and back between the gas syringes. Copper will react with oxygen and undergo oxidation to form Copper (II) oxide. When all the oxygen gas is used up, the volume of air decreases by 21%.



- The oxygen will react with the hot copper and the overall volume of gas in the syringes will drop.



- Volume of air in the gas syringe will decrease by 21% as copper has been oxidized to copper (II) oxide by reacting with the oxygen in the gas syringe. 21% of air contains oxygen gas.
- Volume of gas is read after apparatus cooled to room temperature.
- The experiment is repeated until a constant volume is obtained (to ensure that all the oxygen has been removed.)

Observations

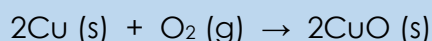
- The **copper** starts off a **reddish-brown colour** and after the reaction is complete it is **black**, the colour of **copper (II) oxide**, CuO.

Application Qns:

200 cm³ of air in syringe (left) was made to pass over heated copper until the reaction was complete. The gas in syringe (right) was then allowed to cool to its original temperature. What is the volume of gas collected in syringe (right)?

Answer:

Copper reacted with oxygen in the air to form copper (II) oxide.

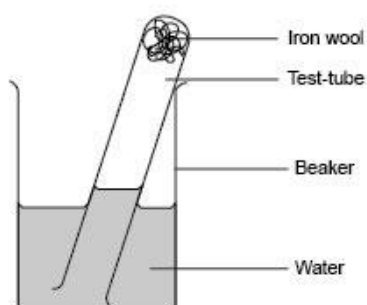
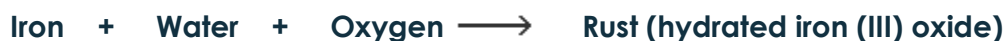


Since oxygen makes up approximately 20% of air, maximum volume of oxygen that will react = 200cm³ x 20% = 40 cm³

Hence, volume of gas collected in syringe B = 200cm³ – 40 cm³ = 160 cm³



Example 2: Rusting



Experimental Setup

Put about 3 cm depth of iron wool into the test tube and wet it with water. Tip away excess water.

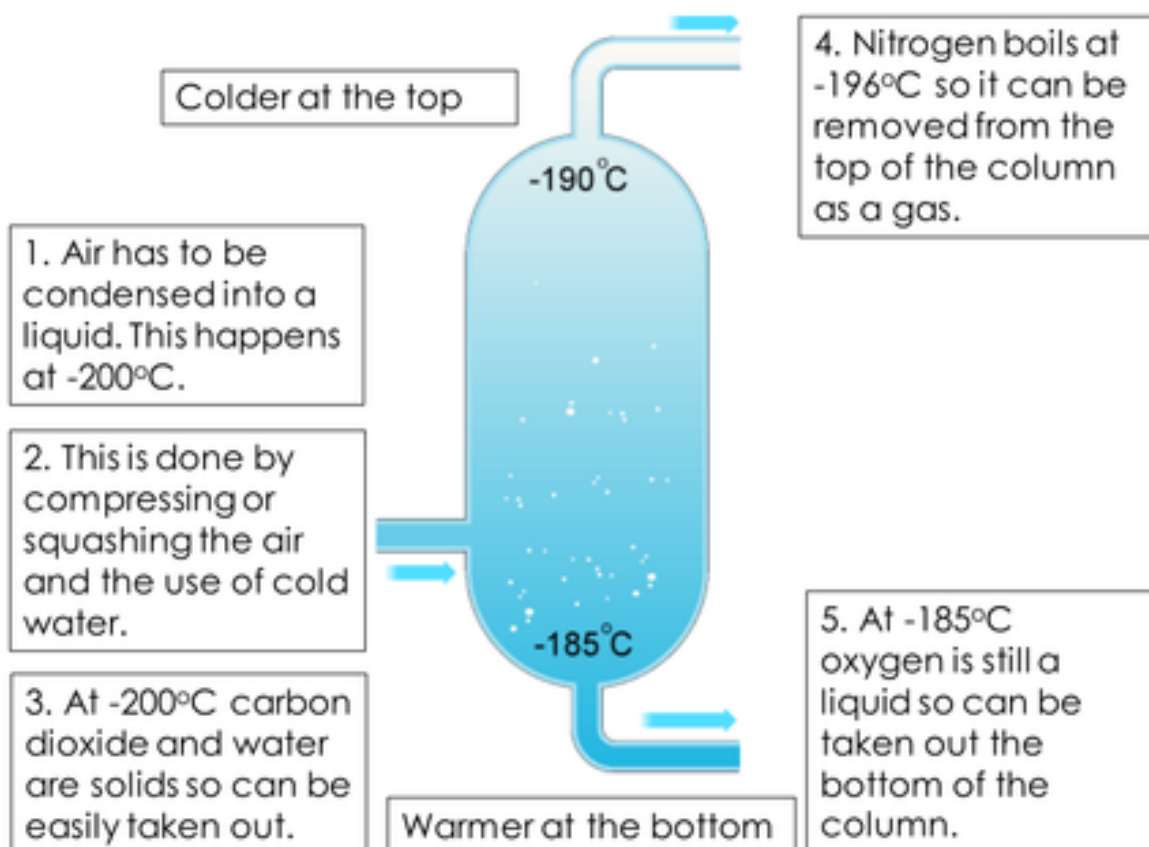
Add water into the beaker. Invert the test tube and place it in the beaker of water (see diagram). Leave for at least a week.

Observation:

The water level would have risen by 21% indicating that 21% of air contains oxygen and oxygen is used in the rusting of iron.

Leave the setup for another 1-2 days and read the water level again. The same reading will ensure that all the oxygen is removed.

Fractional distillation of air





PART III: POLLUTANTS

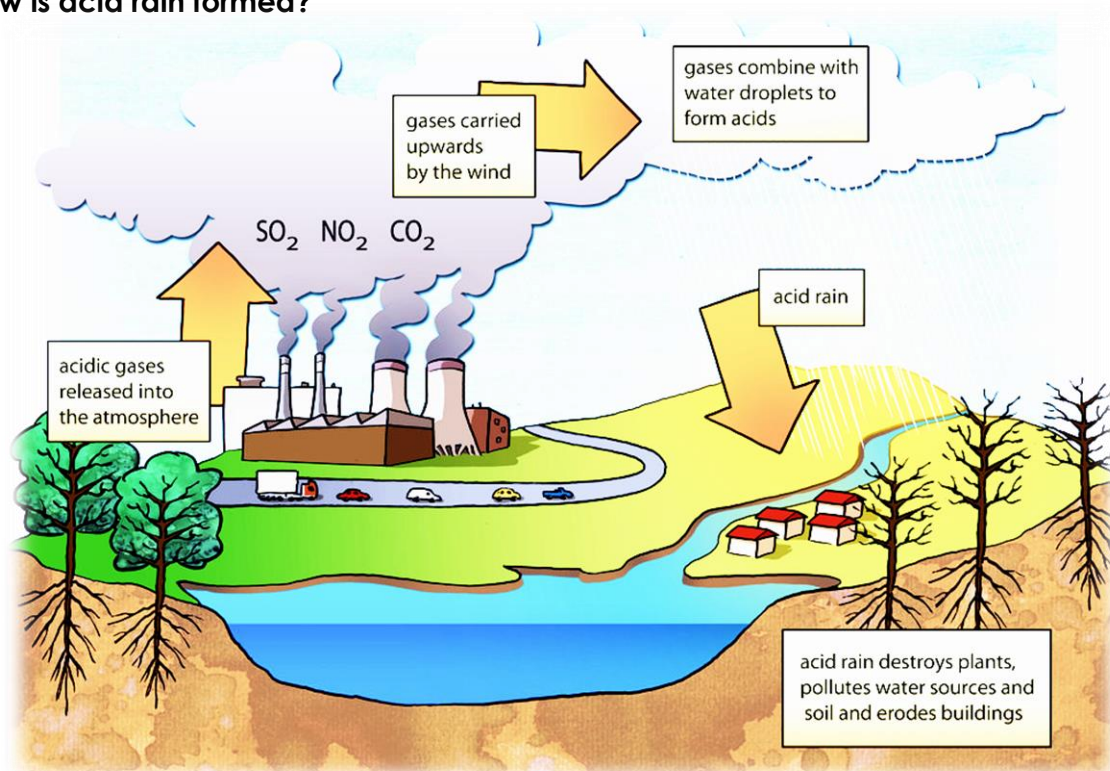
Name of air pollutant	Formula	Type of oxide	Source of pollutant	Problem and effects on environment
Carbon monoxide	CO	Neutral oxide	Due to incomplete combustion (limited oxygen) of carbon containing substances e.g. Charcoal, wood, petrol, etc.	CO combines permanently with haemoglobin in red blood cell and reduce the oxygen intake to the rest of the body. This leads to headaches, fatigue, breathing problems and eventually death.
Oxides of nitrogen	NO nitric oxide or nitrogen monoxide NO ₂ Nitrogen dioxide	NO ₂ Is an Acidic oxide	Internal combustion of car engines Lightning activity In car engines where temperature is high $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}$ Nitrogen monoxide further react to form $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g})$	NO ₂ can dissolve in water to form acid rain . Acid rain corrodes/destroy buildings, aquatic life and plants Irritate eyes and cause breathing difficulties of the lungs
Sulphur dioxide	SO ₂	Acidic oxide	Combustion of fossil fuels in motor vehicles, power stations and factories Gas released during a volcanic eruption $\text{S}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g})$	Acid rain corrodes/destroy buildings. Aquatic life and plants Irritate eyes and cause breathing difficulties of the lungs
Methane	CH ₄		Is a greenhouse gas Formed from decaying organic matter or from decay of rubbish in landfill It is also emitted from the digestion of food by cattle and livestock	Global warming leads to melting of icecaps, rise in sea level.

Note: CO₂ is not a pollutant as it is a gas released by human when we exhale and taken in by plants for photosynthesis. It is the increasing concentration of CO₂ that we are concerned about as it causes global warming.

Definition of pollutant: A substance that causes harm to the environment or pollutes the quality of air, water or land. Eg. Toxic waste dump in water.



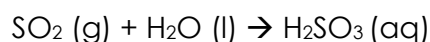
How is acid rain formed?



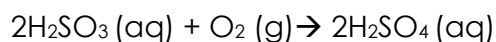
Acid rain is formed when acidic air pollutants such as SO_2 and NO_2 (acidic oxide) react with water in the atmosphere.

- a) Chemical equation for **sulfurous acid, H_2SO_3** (weak acid) formation from sulfur dioxide.

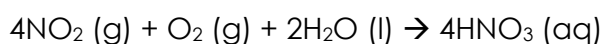
Equation:



*Sulfurous acid, H_2SO_3 can oxidise slowly due to oxygen in air to form **sulfuric acid**, H_2SO_4 . Acid rain causes damages to building and



- b) Nitrogen dioxide reacts with water and oxygen to form nitric acid.





PART IV: Reducing Air Pollution

(a) In Singapore,

- Prohibition of the use of open fires for the disposal of domestic and industrial wastes
- Introduction of unleaded petrol and phasing out of leaded petrol.
(Note: Lead is a particulate pollutant that may cause brain damage.)
- Reduction of the permissible level of sulfur in diesel
- Fitting of all petrol-driven vehicles with catalytic converters

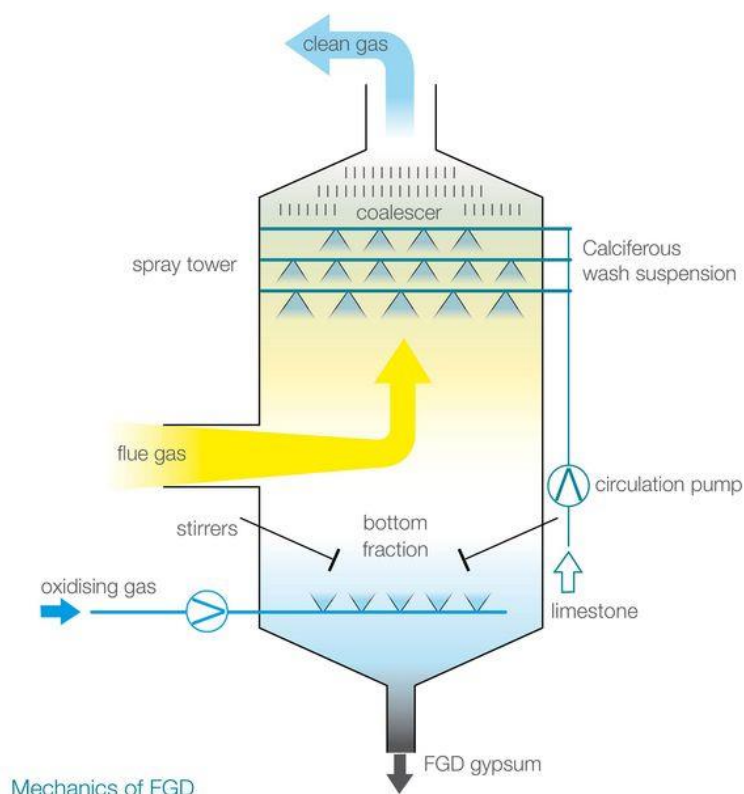
(b) Reducing the effects of acid rain



- Powdered limestone (calcium carbonate) is used to neutralise acid in lakes and rivers. This process is called liming (Shown on the left).
- This method is expensive for farmers and its effects are only temporary.

- Catalytic converters and flue gas desulfurisation (shown on the right) help reduce the emission of acidic gases into the atmosphere.

(Note: **Flue gas**—the emitted material produced when fossil fuels such as coal, oil, natural **gas**, or wood are burned for heat or power—may **contain** pollutants such as particulates, sulfur dioxide, mercury, and carbon dioxide. Most **flue gas**, however, **consists of** nitrogen oxides.)

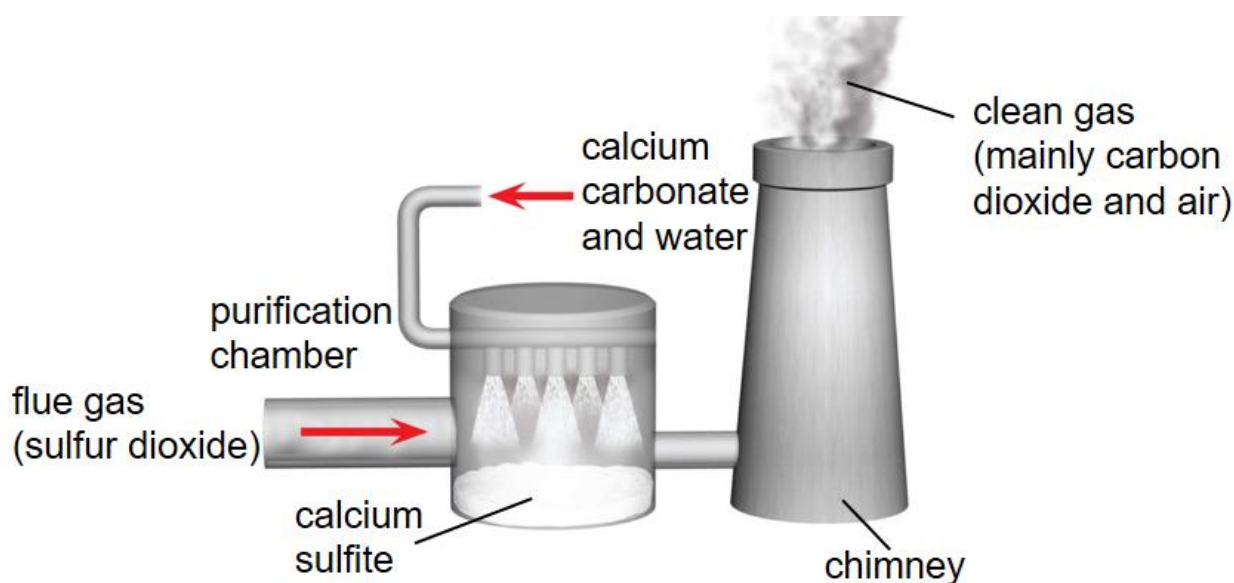




Flue Gas Desulfurisation

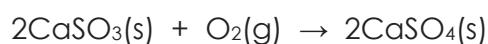
Two ways to minimise the effects of sulfur dioxide pollution:

- Remove sulfur from fossil fuels before burning.
 - However, this method is too expensive and difficult.
- Remove sulfur dioxide from the waste gases formed when the fossil fuels undergo combustion (cheaper).
 - The waste gases are called flue gases.
 - The process of removing **sulfur dioxide** from flue gases is called **desulfurisation**



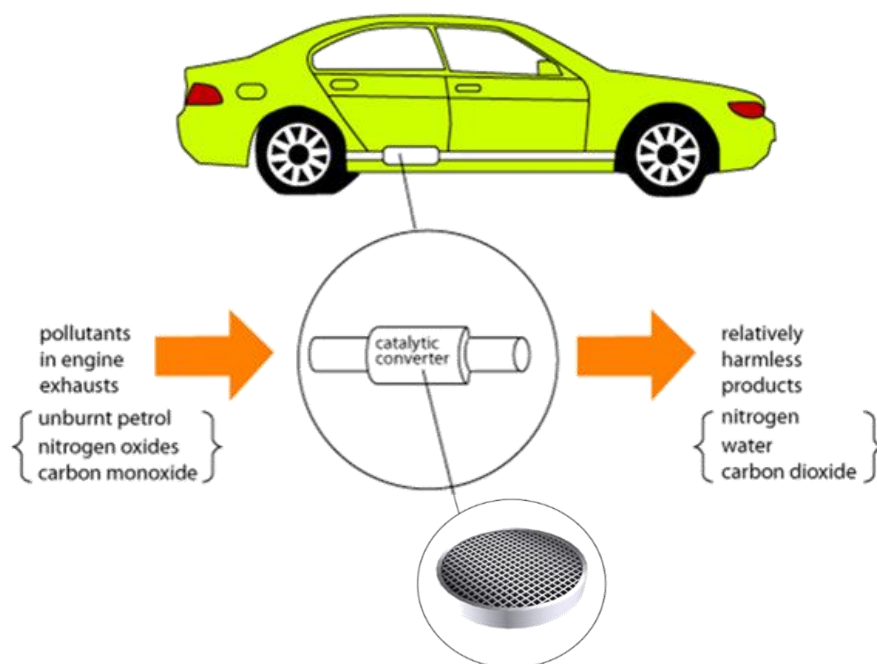
Sulfur dioxide is removed: $\text{CaCO}_3(\text{s}) + \text{SO}_2(\text{g}) \rightarrow \text{CaSO}_3(\text{s}) + \text{CO}_2(\text{g})$

Calcium sulfite is oxidized to calcium sulfate (gypsum, used to make cement):

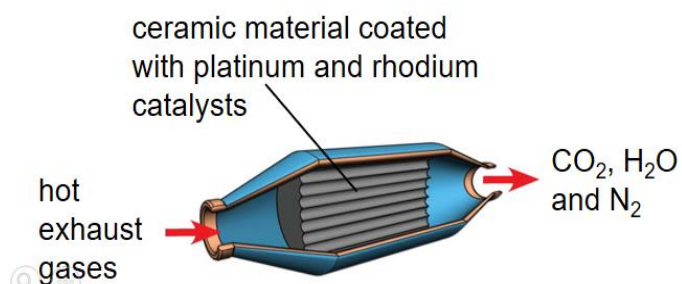




Catalytic Converters



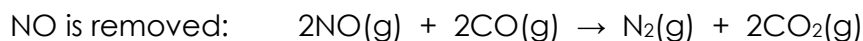
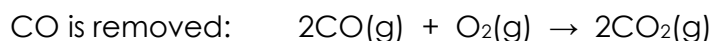
A catalytic converter helps to remove acidic gases in vehicle exhaust gases.



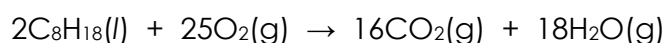
When the hot exhaust gases pass over the platinum and rhodium catalysts, the harmful pollutants undergo redox reactions and are converted into harmless substances.

How pollutants are removed via the catalytic converters?

Carbon monoxide is oxidised to carbon dioxide while nitric oxide is reduced to nitrogen:



Unburnt hydrocarbons are oxidised to carbon dioxide and water:



Facts about the catalytic converters

Why do catalytic converters only work properly after a car has travelled a few kilometres?

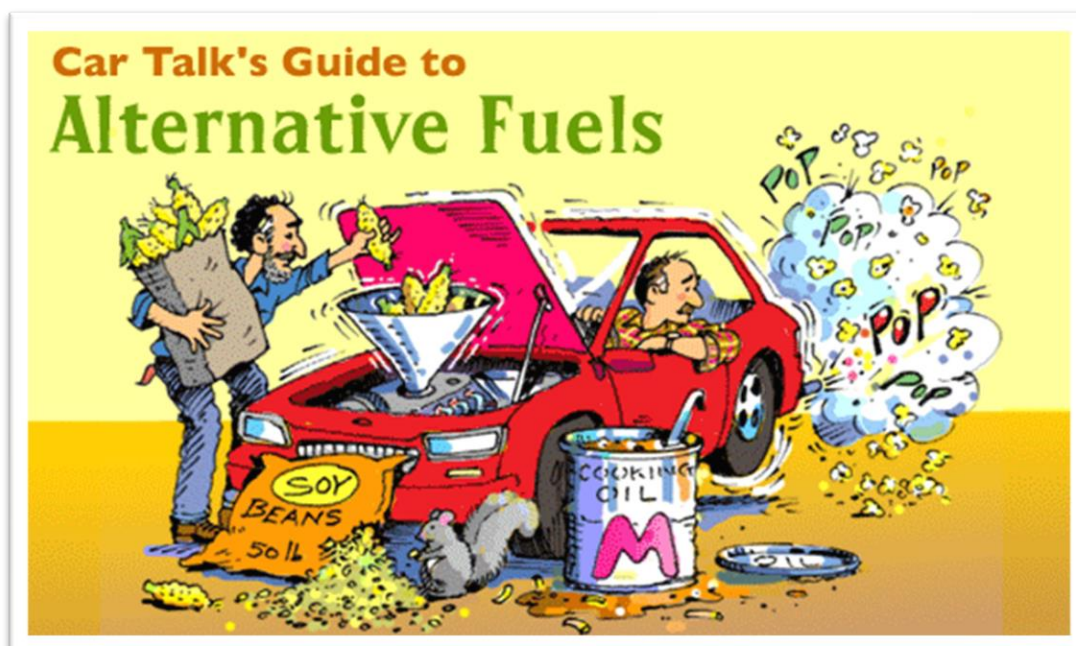
The converter needs to be hot (about 300 °C) for the reactions to be fast enough.

How long can the catalyst last?

Usually last for about 80,000 km until it has to be replaced as the platinum has worn away or stop working.

What happens to the catalytic converter after it no longer functions?

Platinum is expensive, used converters are collected and the platinum are recycled.



Other Measures to Reduce Air Pollution Caused by Motor Vehicles

- Burn less fossil fuels.
- Use lightweight alloys to make car bodies to improve fuel efficiency.
- Switch to electric or hybrid electric vehicles to reduce exhaust gas emissions.
- Use alternative fuels such as hydrogen fuel, where products of combustion are harmless; solar and wind energy.





Ozone (O₃)

Pale blue gas found in the upper layer of the atmosphere called stratosphere (20–50 km above sea level)

Acts as a shield to filter out harmful ultraviolet radiation from the Sun.

Ozone Depletion Problem

- In recent years, the ozone layer is depleting.
- This allows some harmful UV rays to reach the Earth's surface. This can lead to cataract and skin cancer in humans.
- Caused by a group of compounds called chlorofluorocarbons (CFCs)

Chlorofluorocarbons (CFCs)

Contain the elements **carbon**, **fluorine** and **chlorine**

Were widely used as propellants for aerosols and as coolants in refrigerators and air conditioners

Were also used in the manufacture of packing foam

Chlorine atoms in CFCs **break down ozone**

Fluorine gives stability to CFCs, making them persistent in the atmosphere (difficult to remove)



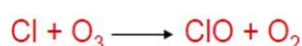
Solution

- In 1992, an international agreement was reached for a complete ban on the release of CFCs by 1996.
- However, even then, the CFCs that are already present in the atmosphere will continue to deplete the ozone layer.

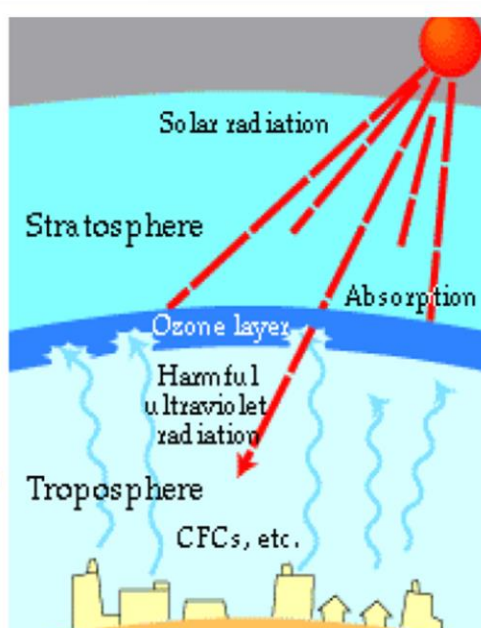
- CFCs reach the stratosphere and the uv rays decomposes the CFC to produce Cl atoms.



- Cl atoms react with ozone to form chlorine oxide and oxygen.



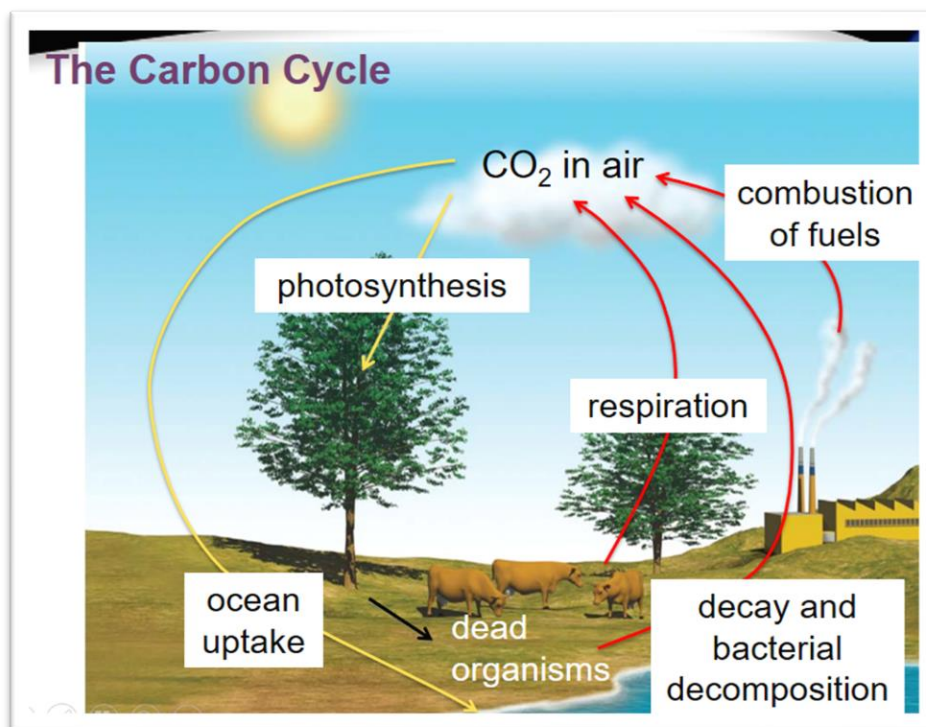
The reaction between Cl atoms and O₃ molecules destroys the ozone layer.





The Carbon Cycle

The Carbon cycle - maintains the level of carbon dioxide in the atmosphere



For the atmosphere to maintain a constant amount of carbon dioxide,

$$\text{Rate of removal of CO}_2 = \text{rate of return of CO}_2$$

- Photosynthesis
- Ocean uptake
- Respiration of animals and man
- Combustion of fuels
- Decay and bacterial decomposition

The following two processes produce carbon dioxide:

Respiration	Combustion of fuels
During respiration, glucose in food we eat is converted into carbon dioxide and water.	The burning of fuels like methane releases carbon dioxide.
Energy is also released. (Exothermic)	Energy is also released. (Exothermic)
$\text{C}_6\text{H}_{12}\text{O}_6(\text{aq}) + 6\text{O}_2(\text{g}) \rightarrow 6\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{l}) + \text{energy}$	$\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g}) + \text{energy}$

The Photosynthesis removes carbon dioxide:

During photosynthesis, green plants convert carbon dioxide and water into glucose and oxygen in the presence of sunlight:





Greenhouse Effect and Global warming

Certain gases in our atmosphere allow visible radiation from the Sun to reach the Earth's surface. However, they trap some of the infrared radiation emitted by the Earth.

This process is called the greenhouse effect.

Gases which trap the energy are called greenhouse gases.

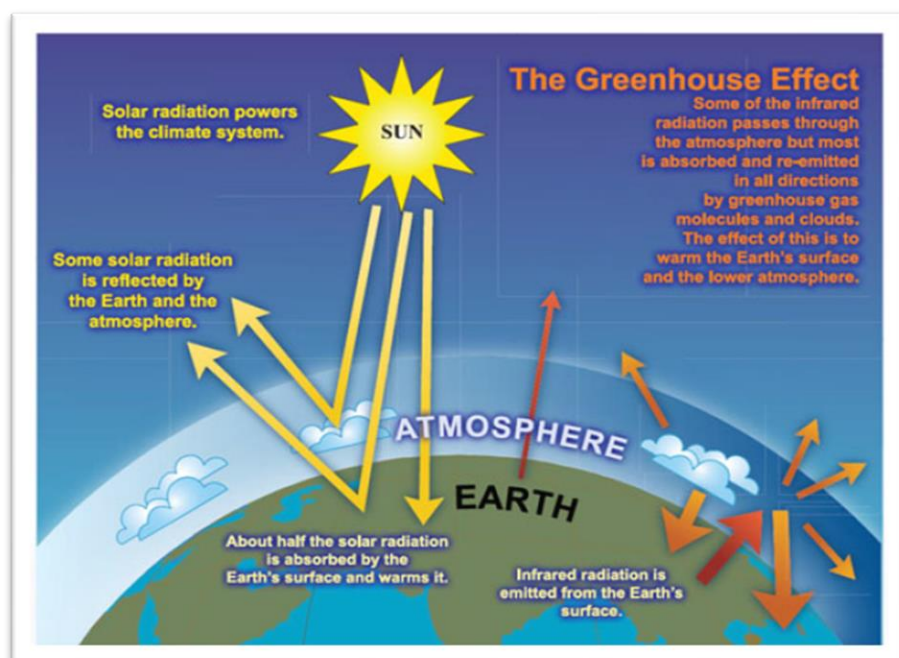
They are mainly carbon dioxide, methane (CH_4) and water vapour.

Without these gases, Earth's surface would have a temperature of about -40°C .

Global Warming

Activities like the burning of fossil fuels and large-scale cutting down of forests are causing some greenhouse gases, especially carbon dioxide, to build up rapidly in the atmosphere.

This means that carbon dioxide is being added to the atmosphere at a higher rate than photosynthesis can remove the excess gas.



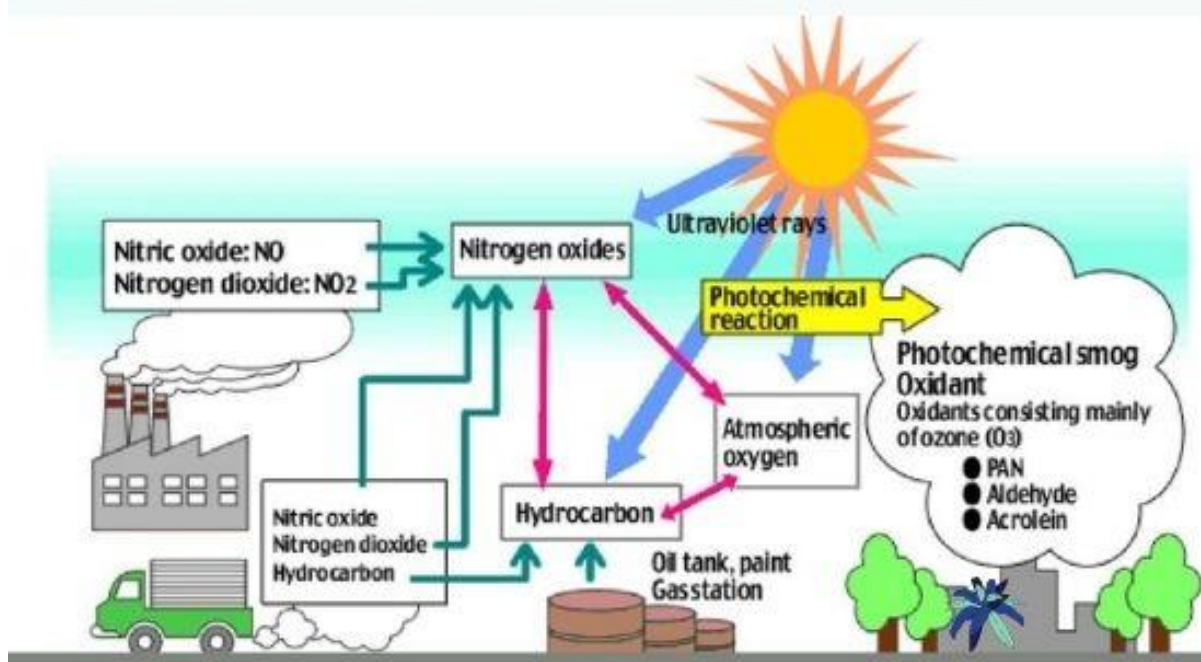
The effect of carbon dioxide build-up leads to an increase in the Earth's average temperature. This is global warming.

Consequences of Global Warming

- Melting of large quantities of ice in the North Pole and South Pole, which may lead to floods
- More occurrences of unusual weather conditions; dry places get drier while wet places get wetter
- Decrease in crop yields world-wide



FORMATION OF PHOTOCHEMICAL SMOG

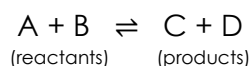




AMMONIA

Part I: Reversible Reaction

A reversible reaction is a chemical reaction where the reactants form products that, in turn, react together to give the reactants back.



Reversible reactions will reach an equilibrium point where the concentrations of the reactants and products will no longer change.

Example:

Reaction of nitrogen and hydrogen to form ammonia is a reversible reaction.

Reversible reaction	comprising of	Forward rxn	$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$
$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$		Backward rxn	$2\text{NH}_3(\text{g}) \rightarrow \text{N}_2(\text{g}) + 3\text{H}_2(\text{g})$

A reversible reaction can go both forward and backward at the same time. A double arrow sign, \rightleftharpoons , is used to indicate a reversible reaction. The reaction from left to right is called the forward reaction. The reaction from right to left is called the backward reaction.

SESD

Ammonia in Fertilizer

About 90 percent of ammonia produced is used in fertilizer, to help sustain food production for billions of people around the world. The production of food crops naturally depletes soil nutrient supplies. In order to maintain healthy crops, farmers rely on fertilizers to keep their soils productive.

Ammonia in Household Cleaning Products

On its own or as an ingredient in many household cleaning products, ammonia can be used to clean a variety of household surfaces – from tubs, sinks and toilets to bathroom and kitchen countertops and tiles.

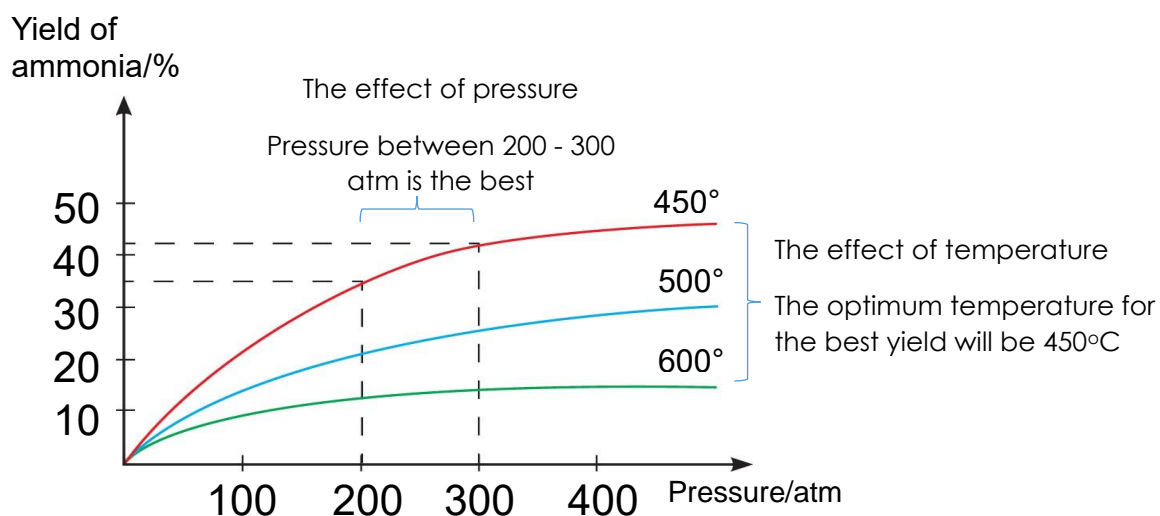
Ammonia in Industrial/Manufacturing Uses

Ammonia also is used in the waste and wastewater treatment, cold storage, rubber, pulp and paper and food and beverage industries as a stabilizer, neutralizer and a source of nitrogen. It also is used in the manufacture of pharmaceuticals.



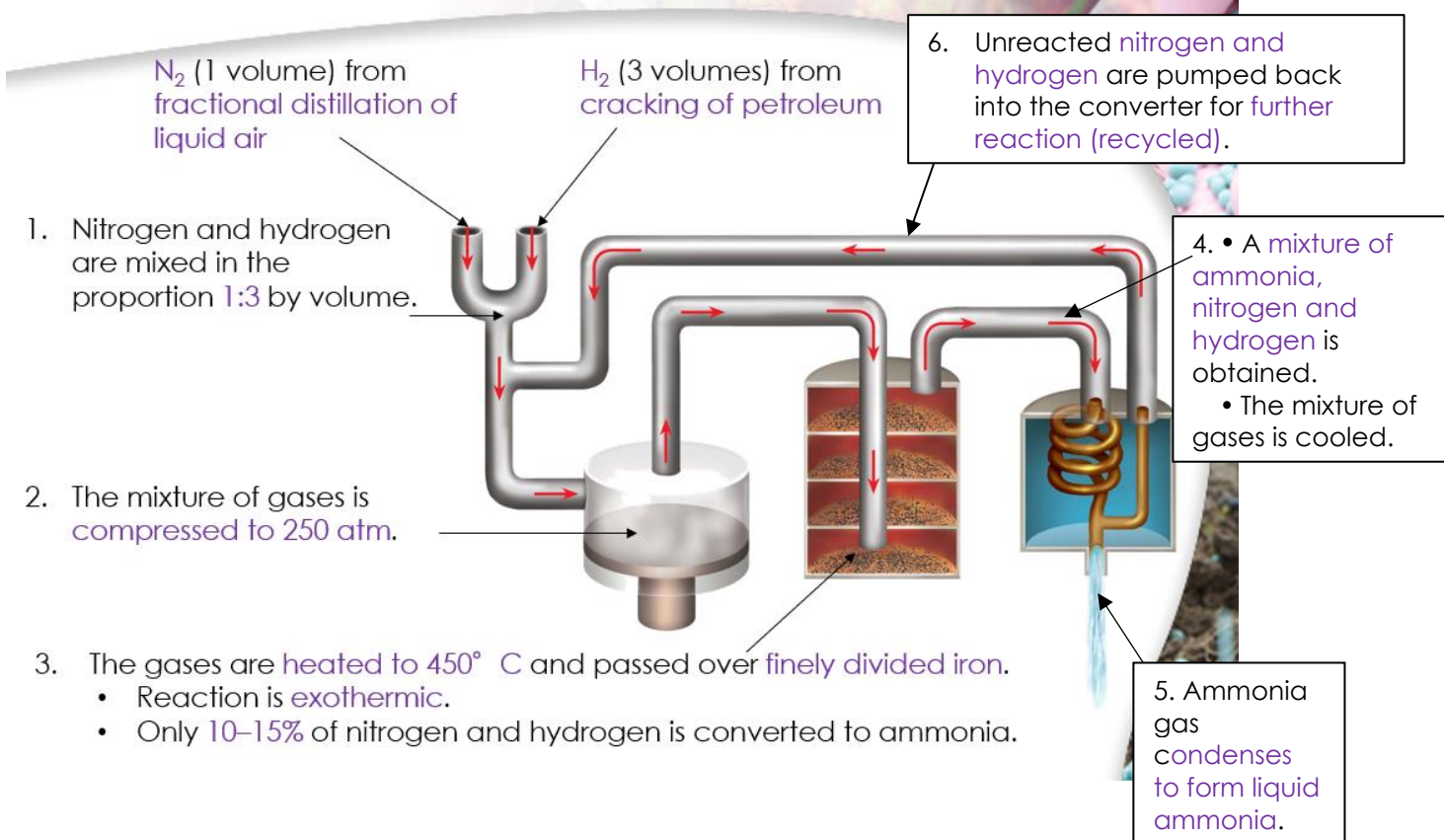
Part II: Haber process (Industrial preparation of Ammonia)

Raw materials	Nitrogen gas	(from fractional distillation of liquid air)
	Hydrogen gas	(from cracking of petroleum)
Conditions	High pressure 250 atm	<p>The higher the pressure, the higher the yield of ammonia. High pressure also increases the speed of the reaction.</p> <p>Maintaining high pressure is costly.</p> <p>Thus, there is a limit to the amount of pressure that can be applied.</p>
	High temperature 450°C	<p>The lower the temperature, the higher the yield of ammonia. This is because the decomposition of ammonia into H₂ and N₂ is reduced.</p> <p>Lower temperature also results in slower reaction.</p> <p>Thus, a relative high temperature of 450°C is used.</p>
Catalyst	Finely divided Iron	
Equation	$\text{nitrogen} + \text{hydrogen} \xrightleftharpoons{\text{finely divided iron}} \text{ammonia}$	





Manufacturing Ammonia by the Haber Process



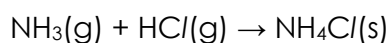
Part III: Reactions of Ammonia & its compounds

1. Thermal decomposition of ammonium chloride

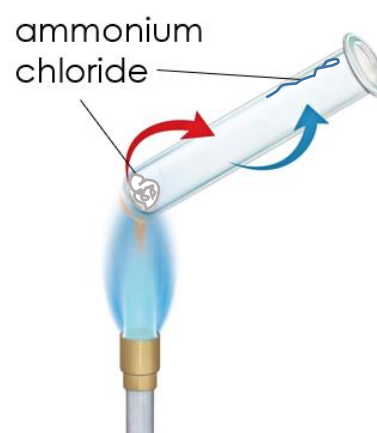
Solid ammonium chloride decomposes upon heating to form ammonia gas and hydrogen chloride gas.



Upon cooling, solid ammonium chloride reforms.



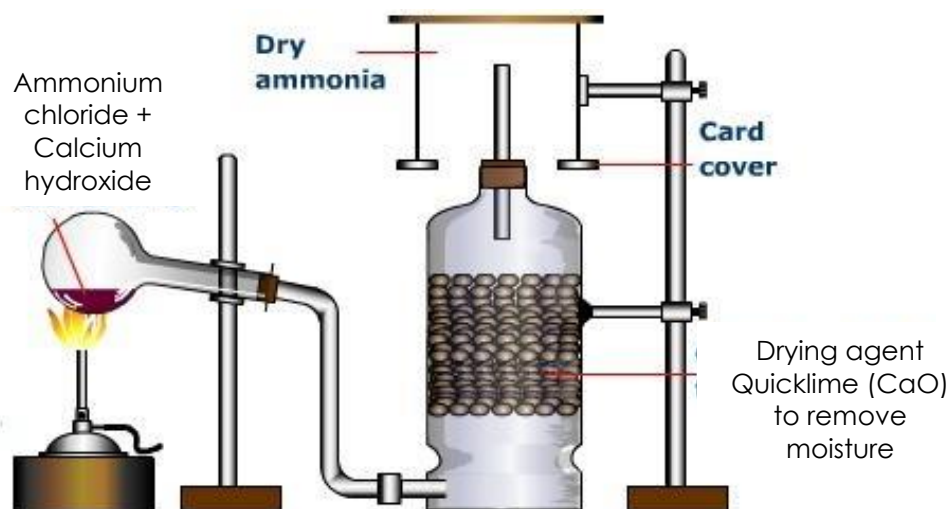
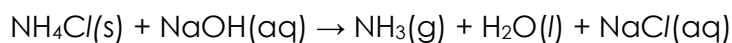
Reversible reaction: $\text{NH}_4\text{Cl(s)} \rightleftharpoons \text{NH}_3\text{(g)} + \text{HCl(g)}$





2. Displacement of Ammonia from its Salts

Whenever an ammonium salt is heated with an alkali, ammonia is displaced from the salt.



Ammonium salts	Alkali
Ammonium nitrate	Sodium hydroxide
Ammonium chloride	Potassium hydroxide
Ammonium sulfate	

Addition of calcium hydroxide (to neutralize acidic solid) together with fertilizers (containing ammonium salts) will result in the formation of ammonia gas. This will remove the element nitrogen which is important for plant growth.

3. Cations Test (QA)

Cations	Observation
Ca^{2+}	No ppt formed
Al^{3+}	White ppt insoluble in excess NH_3 (aq)
Pb^{2+}	White ppt insoluble in excess NH_3 (aq)
Zn^{2+}	White ppt soluble in excess NH_3 (aq) to form a colourless solution
Cu^{2+}	Blue ppt soluble in excess NH_3 (aq) to form a dark blue solution
Fe^{2+}	Green ppt insoluble in excess NH_3 (aq)
Fe^{3+}	Brown ppt insoluble in excess NH_3 (aq)



Organic Chem - Fuel

Introduction - What are Organic compounds?

- All organic compounds contain carbon.
- Most organic compounds also contain hydrogen.
- Hydrocarbons are organic compounds that contain only hydrogen and carbon.
- Some organic compounds may also contain oxygen, chlorine and nitrogen.

Homologous Series

- To study the different organic compounds, related organic compounds are grouped into families called homologous series.
- A **homologous series** is a family of organic compounds with the same functional group and similar chemical properties.
- Examples of homologous series:
 - alkanes
 - alkenes
 - alcohols
 - carboxylic acids

Functional Group

Homologous Series	Functional group	Structural representation
Alkanes	No functional group	
Alkenes	Carbon- Carbon double bond	$\begin{array}{c} \diagup \quad \diagdown \\ \text{C} = \text{C} \\ \diagdown \quad \diagup \end{array}$
Alcohol	Hydroxyl group	$-\text{O}-\text{H}$
Carboxylic acids	Carboxyl group	$\begin{array}{c} \text{O} \\ \\ -\text{C}-\text{O}-\text{H} \end{array}$



Organic compounds in the same homologous series

- have the same functional group;
- have similar chemical properties;
- show a gradual change in their physical properties down the series.

Naming Organic Compounds

The name of an organic compound is divided into two parts:

1. **Prefix** — shows the number of carbon atoms in each molecule.

No. of carbon atoms per molecule	Prefix	No. of carbon atoms per molecule	Prefix
1	Meth ~	5	Pent ~
2	Eth ~	6	Hex ~
3	Prop ~	7	Hept ~
4	But ~	8	Oct ~

2. **Suffix** — shows the homologous series that the compound belongs to.

Homologous series	Suffix	Homologous series	Suffix
Alkane	~ ane	Alcohol	~ anol
Alkene	~ ene	Carboxylic acid	~ anoic acid

Example

No. of carbon atoms	Alkane	Alkene	Alcohol	Carboxylic acid
1	Methane	-	Methanol	Methanoic acid
3	Propane	Propene	Propanol	Propanoic acid
5	pentane	Pentene	Pentanol	Pentanoic acid

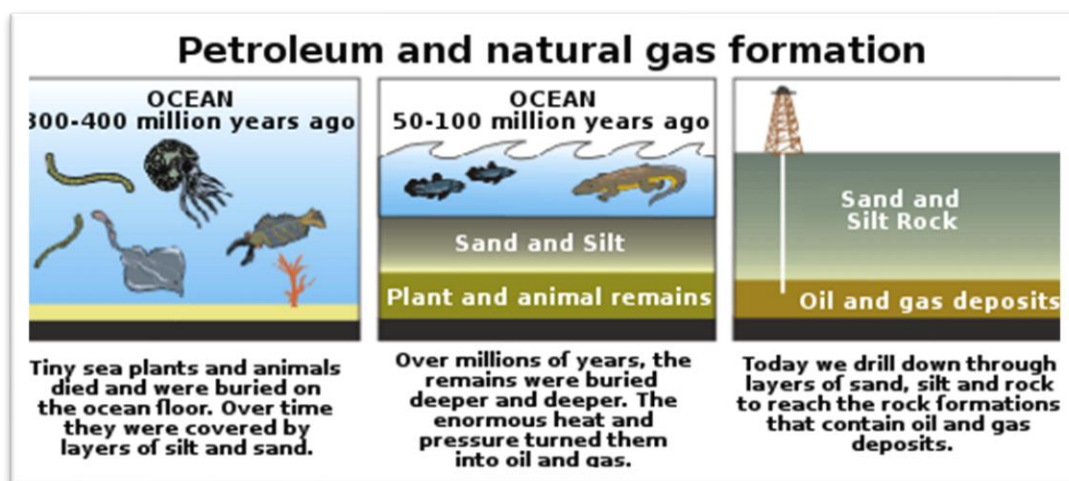


Petroleum and Natural Gas

- **Petroleum** and **natural gas** are used as sources of energy.
- **Petroleum** is a naturally occurring mixture of hydrocarbons (mainly alkanes).
- Petroleum is also called crude oil.
- **Natural gas** is mostly made up of
 - methane (about 70–90%);
 - other short-chain alkanes such as ethane (about 0–20%), propane and butane.

How are Petroleum and Natural gas formed?

- Sea creatures and plants sink to the seabed when they die.
- Over millions of years, these organisms are subjected to heat from the Earth and pressure.



- Petroleum and natural gas are formed.
- Deep wells are drilled to get them out.

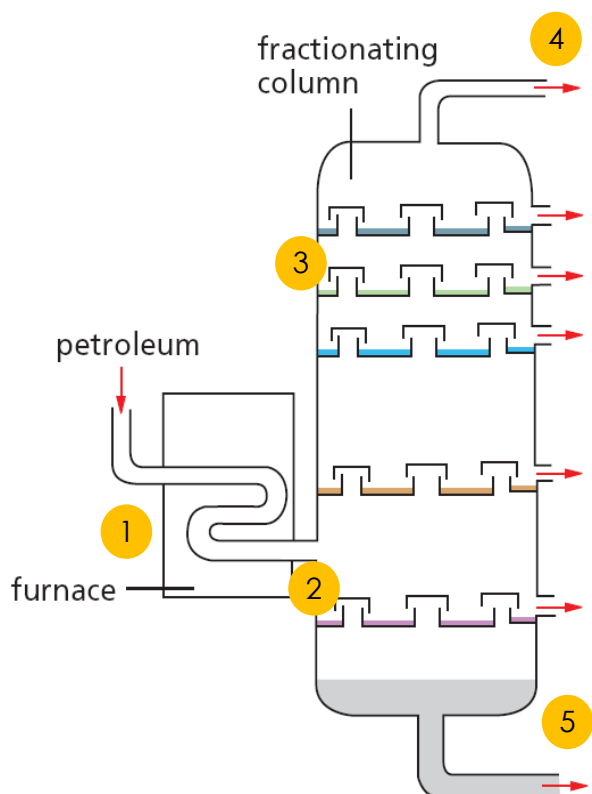
Fractional Distillation of Petroleum

Petroleum can be separated into useful fractions by **fractional distillation**.

- Each fraction is a mixture of hydrocarbons which boils over a certain temperature range.
- A lighter fraction contains fewer carbon atoms and has a lower boiling point range than a heavier fraction.

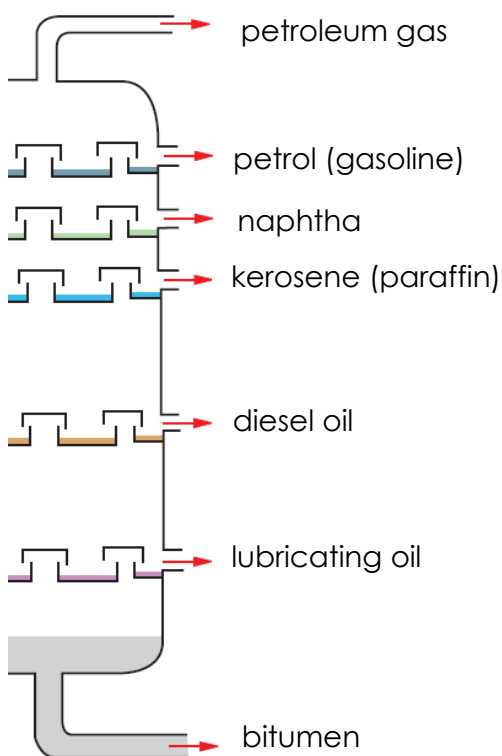


How is Petroleum Separated in an Oil Refinery?



Step	Process
1	Petroleum is heated into a vapour in the furnace.
2	The vapour is passed into the fractionating column.
3	The hot vapour cools and condenses as it rises up the column
4	Lighter fractions with lower boiling points are collected at the top of the column as gases.
5	Heavier fractions with higher boiling points are collected at the bottom of the column as residue.

Uses of Petroleum Fractions



Fractions	Process
Petroleum gas	fuel for cooking and heating
Petrol	fuel for car engines
Naptha	Feedstock (raw material) for petroleum industry (which produces plastics, detergents etc)
Kerosene	fuel for aircraft engines, cooking and heating
Diesel Oil	fuel for diesel engines in buses, lorries and trains
Lubricating Oil	for lubricating machines; for making waxes and polishes
Bitumen	for paving road surfaces; for roofing



Issues Related to the Competing Uses of Petroleum

- Petroleum is a non-renewable resource. This means that supply of petroleum is limited or finite
- To meet the two competing uses (chemical feedstock for plastics etc and fuel) of petroleum, there is a growing need for its conservation.

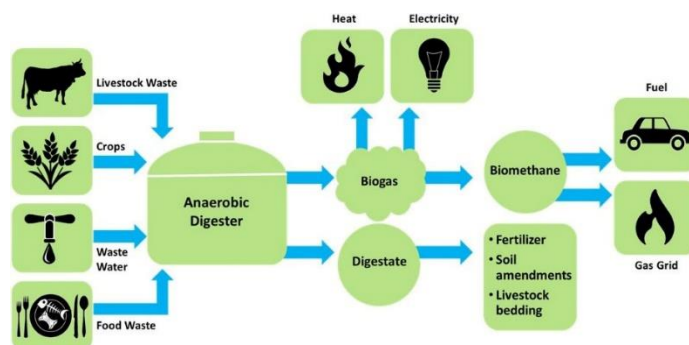
Ways to Conserve Petroleum

- Reduce the number of motor vehicles on the road.
- Take public transport such as buses and the MRT
- Drive cars that consume less petrol
- Improve the design of power stations and vehicles for more efficient use of petroleum.
- Use alternative energy sources such as solar energy and nuclear energy.
- Use alternative fuels.

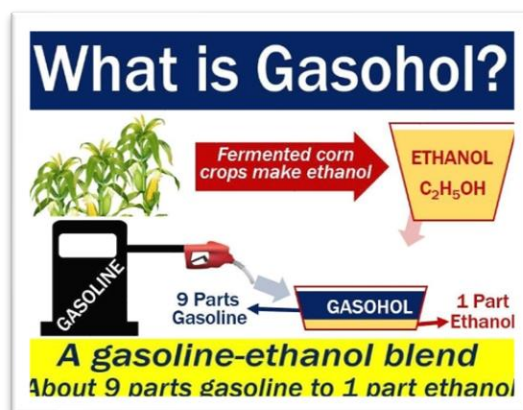


Alternative Fuels

- Biofuel



- biogas (produced when dead plants or animals decay in the absence of air)
- biodiesel (made from plant oils like palm oil)
- Gasohol
 - a mixture of ethanol and petrol





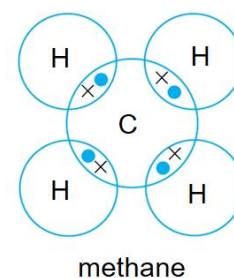
Alkanes

Alkanes are hydrocarbons that contain only carbon–carbon single covalent bonds

- Alkanes have the general molecular formula C_nH_{2n+2} (where $n = 1, 2, 3$ etc.).
- Their names end with '–ane'.
- Each member of the alkane homologous series differs from the next by a $-CH_2-$ unit.

No. of C atoms	Prefix	Name (~ane)	Molecular formula C_nH_{2n+2} ($n = \text{C atoms}$)	Full Structural formula
1	Meth-	Methane	$C_1H_{2(1)+2} = CH_4$	<pre> H H — C — H H </pre>
2	Eth-	Ethane	C_2H_6	<pre> H H H — C — C — H H H </pre>
3	Prop-	Propane	C_3H_8	<pre> H H H H — C — C — C — H H H H </pre>

- Alkanes are saturated hydrocarbons.
- They contain *only* carbon–carbon single covalent bonds.
- For example, in methane, CH_4 , all the valence electrons of the carbon atom are used to form single bonds with four hydrogen atoms.





Isomerism in Alkanes

Isomers have the same molecular formula but different structural formulae.

- Isomers have different physical properties such as melting and boiling points.
- E.g. butane (b.p.: -0.5°C) and methylpropane (b.p.: -11.7°C)
- They have the same molecular formula, C_4H_{10} , but different structural formulae.

Butane	Methyl - propane
<pre> H H H H H - C - C - C - C - H H H H H </pre>	<pre> H H H H - C - C - C - H H H H-C-H H </pre>

Step 1: Name the longest C Chain

Step 2: Look any branched out chain or alkyl group.

Step 3: Count which C atom on the main chain the alkyl group is attached.

- Alkyl groups** have the general molecular formula $\text{C}_n\text{H}_{2n+1}$.

- Examples of alkyl groups:

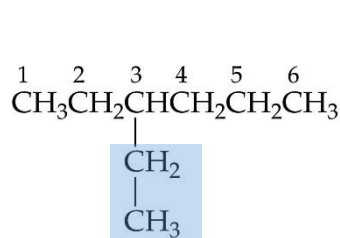
methyl: CH_3-

ethyl: CH_3CH_2-

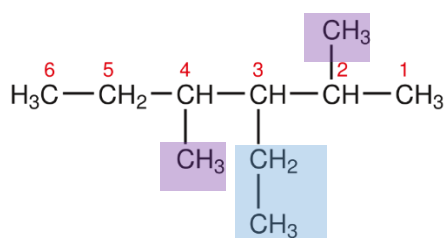
propyl: $\text{CH}_3\text{CH}_2\text{CH}_2-$

butyl: $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2-$

Examples



3 ethyl hexane



3-ethyl-2,4-dimethylhexane

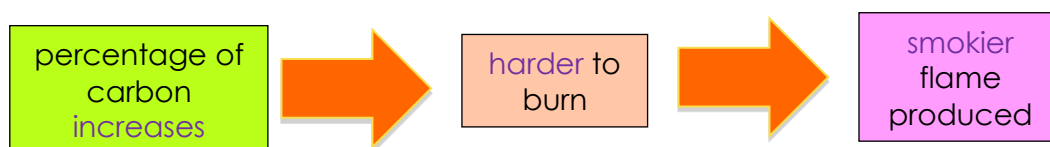


Physical Properties of Alkanes

- Low melting and boiling points
- Insoluble in water
- Soluble in organic solvents
- Going down the alkane homologous series, molecular size of alkanes increases; intermolecular forces of attraction between alkane molecules increase; more energy is required to overcome the intermolecular forces between the molecules.

Thus,

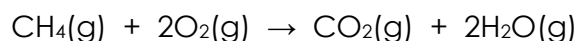
- the melting and boiling points of alkanes increase down the homologous series; (Larger molecules have increased intermolecular forces of attraction)
- alkanes become more viscous i.e. flow less easily;
- their flammability decreases.



Chemical Properties of Alkanes

1. Complete Combustion

- Alkanes burn in excess air (oxygen) to produce carbon dioxide and water vapour.



- This reaction is highly exothermic (produces a lot of heat), hence alkanes are used as fuels.

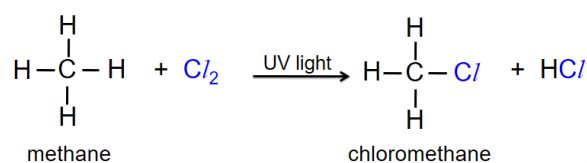
2. Incomplete Combustion

When air (oxygen) is limited, alkanes will burn to produce carbon monoxide (toxic), soot and water.

3. Substitution reactions

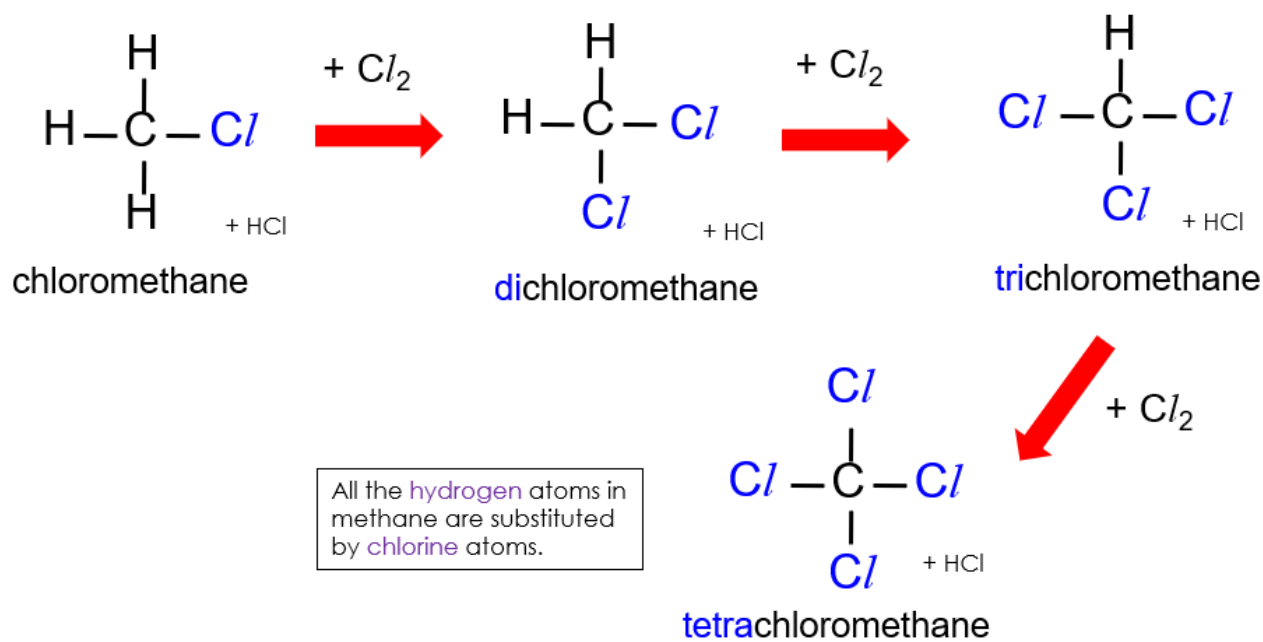
Alkanes react with halogens (e.g. Cl_2 , Br_2) in the presence of ultraviolet (UV) light.

Reaction of methane with chlorine





Complete reaction of methane and chlorine





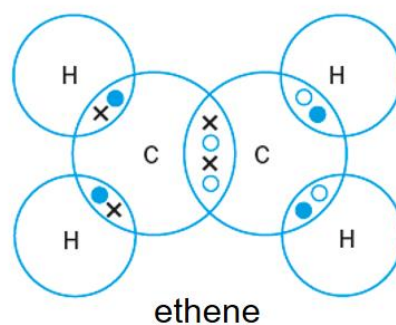
Alkenes

Alkenes are hydrocarbons that contain one or more carbon–carbon double bonds.

- Alkenes have the general molecular formula C_nH_{2n} (where $n = 2, 3, 4$, etc.).
- They contain the $C=C$ functional group.
- Their names end with '–ene'.
- Each member of the alkene homologous series differs from the next by a $-CH_2-$ unit.

No. of C atoms	Prefix	Name (~ene)	Molecular formula C_nH_{2n} ($n = \text{C atoms}$)	Full Structural formula
2	Eth-	Ethene	C_2H_4	<pre> H H \ / C = C / \ H H </pre>
3	Prop-	Propene	C_3H_6	<pre> H H H H - C - C = C - H H </pre>
4	But-	Butene	C_4H_8	<pre> H H H H H - C - C - C = C - H H H </pre>

- Alkenes are unsaturated hydrocarbons.
- They contain carbon–carbon double covalent bonds
- For example, in ethene, C_2H_4 , a $C=C$ bond is formed by the sharing of two pairs of electrons between the carbon atoms.





Isomerism in Alkenes

For example, butene and methylpropene are isomers.

- They have the same molecular formula C_4H_8 but different structural formulae.
- They have different physical properties such as melting and boiling points.

Butene, C_4H_8	Methylpropene, C_4H_8
$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H} - \text{C} - \text{C} = \text{C} - \text{H} \\ \\ \text{H} \end{array} $	$ \begin{array}{c} \text{H} \qquad \qquad \text{H} \\ \qquad \qquad \\ \text{H} - \text{C} - \text{C} = \text{C} - \text{H} \\ \quad \quad \\ \text{H} \quad \text{H} - \text{C} - \text{H} \\ \\ \text{H} \end{array} $

Physical Properties of Alkenes

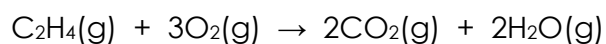
- As molecular size of alkenes increase down the homologous series, the boiling points of alkenes increase. Alkenes are covalent compounds and therefore have weak intermolecular forces between molecules.

However, as the size of the alkene molecules increases, the intermolecular forces increase. Hence, more energy is required to overcome these forces and the boiling points increase.

Chemical Properties of Alkenes

1. Complete Combustion

- Alkenes burn in excess air (oxygen) to produce carbon dioxide and water vapour.



- Alkenes contain a relatively higher percentage of carbon than the corresponding alkanes. Hence, they burn with a sootier flame.

2. Incomplete Combustion

When air (oxygen) is limited, alkanes will burn to produce carbon monoxide (toxic), soot and water.



3. Addition Reactions

- Carbon-carbon double bonds (C=C) in alkenes are very reactivity
- Thus, alkenes readily undergo addition reactions.
- In an **addition** reaction, an unsaturated organic compound combines with another substance to form a single new compound.



Addition Reactions of Alkenes	Equation
Hydrogenation <ul style="list-style-type: none"> The addition of hydrogen to alkenes to form alkanes Conditions: 200°C, nickel catalyst 	
Bromination <ul style="list-style-type: none"> The addition of bromine to alkenes Also used to differentiate alkanes and alkenes based on colour change 	
Hydration <ul style="list-style-type: none"> The addition of steam to alkenes to form alcohols Conditions: 300°C, 60 atm, phosphoric(V) acid catalyst (H₃PO₄) 	
Addition polymerisation <ul style="list-style-type: none"> Alkene molecules (monomers) react with each other to form a long molecule called a polymer. Conditions: high temperature and pressure, catalyst 	



Test for Saturated and Unsaturated hydrocarbons

Bromination

- Used to test for the presence of unsaturated hydrocarbons
- Can be used to distinguish between an alkane and an alkene

Alkanes (saturated)	Alkenes (unsaturated)
<ul style="list-style-type: none"> • Alkanes (eg. Ethane) does not decolourise bromine solution under normal conditions. 	<ul style="list-style-type: none"> • Alkenes (Eg.Ethene) decolourises reddish-brown bromine solution.
<ul style="list-style-type: none"> • Bromine remains reddish-brown 	<ul style="list-style-type: none"> • Bromine solution turns from reddish-brown to colourless.

Production of Alkenes by Cracking

- Cracking is the breaking down of long-chain hydrocarbons into smaller hydrocarbons or hydrogen.
- Alkenes are obtained by cracking petroleum.

Example:

Hexane \rightarrow Butane + Ethene



- In catalytic cracking, a catalyst is used to speed up the process of cracking.

Industrial Catalytic Cracking

For example, the naphtha fraction in petroleum undergoes cracking to produce short-chain alkanes such as ethene and propene, which are used to produce petrochemicals.

Conditions:

- High temperature, e.g. 600°C
- Aluminium oxide (Al_2O_3) and silicon dioxide (SiO_2) catalysts



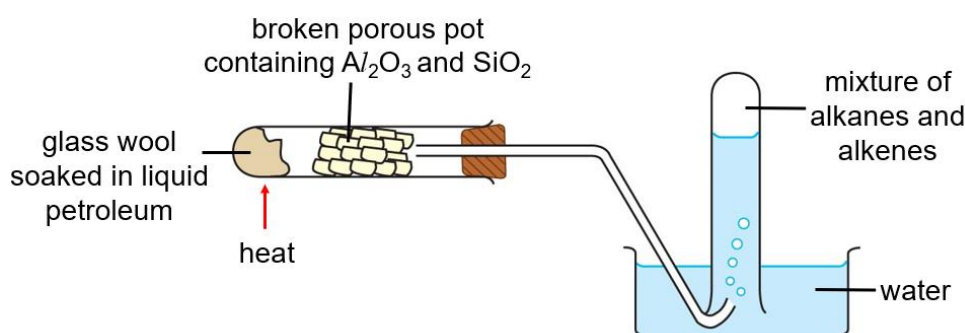
- Modern cracking uses zeolites as catalysts. These are complex aluminosilicates.



Cracking of Hydrocarbons in the School Laboratory

Example: Cracking of liquid petroleum

The set-up below shows how catalytic cracking of liquid petroleum can be carried out in the school laboratory.



Importance of Cracking

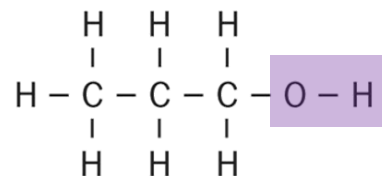
No	Examples
1	<p>Cracking produces short-chain alkenes.</p> <p>Example $\text{C}_{18}\text{H}_{38} \rightarrow \text{C}_6\text{H}_{14} + 6\text{C}_2\text{H}_4$</p> <p>Short-chain alkenes such as ethene and propene are used to make ethanol and plastics.</p>
2	<p>Cracking produces hydrogen.</p> <p>Example $\text{C}_{18}\text{H}_{38} \rightarrow \text{C}_8\text{H}_{16} + \text{C}_{10}\text{H}_{20} + \text{H}_2$</p> <p>Hydrogen is used in the manufacture of ammonia which is used to make fertilisers</p>
3	<p>Cracking converts hydrocarbons of higher molecular mass into smaller molecules which are in higher demand.</p> <ul style="list-style-type: none"> For example, there is a higher demand for petrol than lubricating oil or diesel oil. Petrol can be obtained through catalytic cracking of diesel and lubricating oils.





Alcohols

- **Alcohols** are a homologous series of organic compounds which have the hydroxyl (-OH) functional group. The hydroxyl (-OH) group in alcohols is different from the OH^- ion in alkalis. An alcohol is not an alkali. It is also neutral.



- Alcohols contain carbon, hydrogen and oxygen.
- Alcohols have the general molecular formula $\text{C}_n\text{H}_{2n+1}\text{OH}$ (where $n = 1, 2, 3$ etc.).
- Their names end with '–anol'.
- Each member differs from the next by a $\text{-CH}_2\text{-}$ unit.

No. of C atoms	Prefix	Name (~anol)	Molecular formula $\text{C}_n\text{H}_{2n+1}\text{OH}$ ($n = \text{C atoms}$)	Full Structural formula
1	Meth-	Methanol	CH_3OH	$ \begin{array}{c} \text{H} \\ \\ \text{H} - \text{C} - \text{O} - \text{H} \\ \\ \text{H} \end{array} $
2	Eth-	Ethanol	$\text{C}_2\text{H}_5\text{OH}$	$ \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H} - \text{C} - \text{C} - \text{O} - \text{H} \\ \quad \\ \text{H} \quad \text{H} \end{array} $
3	Prop-	Propanol	$\text{C}_3\text{H}_7\text{OH}$	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{O} - \text{H} \\ \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \end{array} $



Physical Properties

- Alcohols are volatile liquids at room temperature and pressure.

Alcohol	State	Solubility in water	Boiling point /°C
Methanol	Liquid	Very soluble	65
Ethanol	Liquid	Very soluble	78
Propanol	Liquid	Soluble	97
Butanol	Liquid	Slightly soluble	118

Additional Notes

- Due to the highly polar nature of the –OH group in alcohols, stronger hydrogen bonding occurs between alcohol molecules in addition to the weak intermolecular forces (van der Waal forces). More energy is thus required to separate the alcohol molecules, resulting in higher boiling points. Since the boiling points of alcohols are above room temperature, they exist in the liquid state at room temperature.

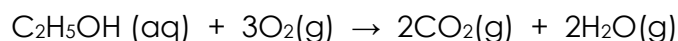
Students are not required by the syllabus to understand polarity of molecules nor hydrogen bonding.

- As the molecular size of alcohols increases,
 - solubility of alcohols decreases; Short-chain alcohols dissolve easily in water due to the presence of the –OH group. However, as the hydrocarbon chain becomes longer, the alcohols become more non-polar and become less soluble in water. Thus solubility of alcohols decreases down the series.
 - boiling point of alcohols increases. This is because the forces of attraction between the alcohol molecules increase.

Chemical Properties

1. Complete Combustion

- Alcohols burn in excess air (oxygen) to produce carbon dioxide and water vapour.



This reaction is highly exothermic (produces a lot of heat), hence alcohols are used as fuels.



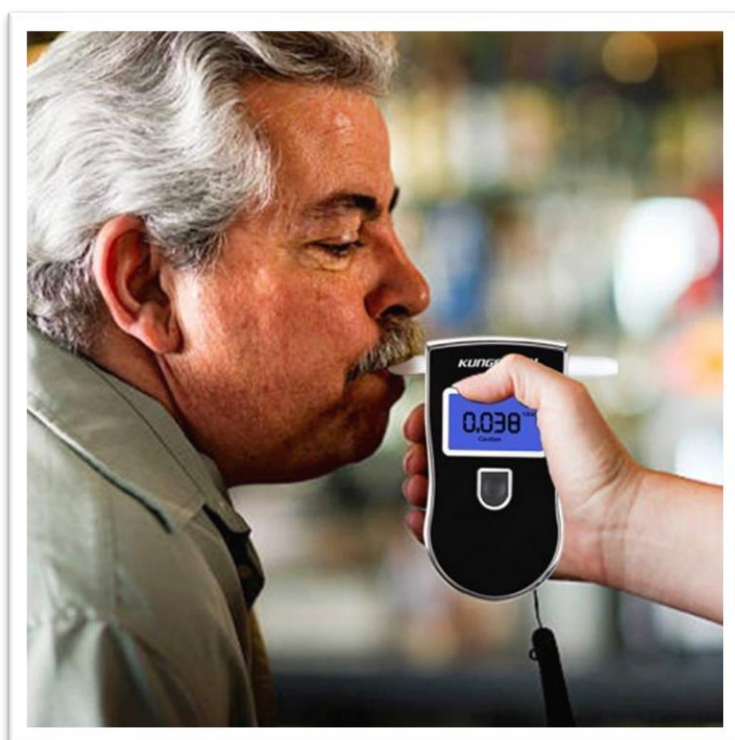
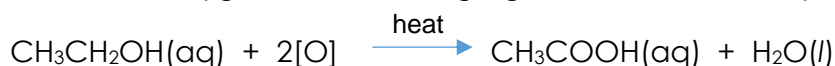
2. Incomplete Combustion

When air (oxygen) is limited, alcohol will burn to produce carbon monoxide (toxic), soot and water.

3. Oxidation

Alcohols are oxidised to carboxylic acids when heated with oxidising agents such as acidified potassium manganate(VII).

alcohol + oxygen from oxidising agent $\xrightarrow{\text{heat}}$ carboxylic acid + water



Breathalyzers

- The police use breathalyzers to test the amount of alcohol consumed by drivers.
- A breathalyser contains an oxidising agent such as acidified potassium manganate(VII).
- A colour change is registered if a certain amount of alcohol is present in the breath.

Producing Ethanol

1. Producing Ethanol by Fermentation

Ethanol can be produced from the fermentation of carbohydrates in fruits, vegetables or grain.

Alcohol fermentation is a process in which micro-organisms such as yeast act on carbohydrates such as glucose in the absence of oxygen to produce ethanol and carbon dioxide.

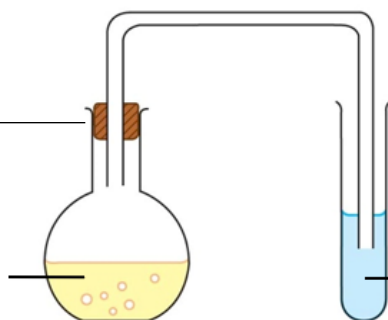




Preparation of Ethanol in the Laboratory

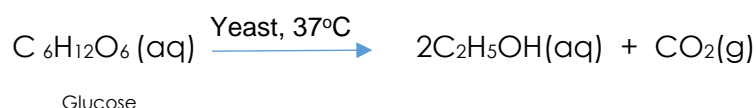
Fermentation takes place only in the absence of oxygen. Hence the flask must be tightly stoppered.

glucose
solution +
yeast



The limewater
acts as an air
lock.

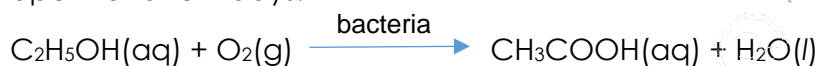
limewater



- Glucose solution is mixed with yeast and kept at 37°C and oxygen must not be present
- Carbon dioxide is produced during fermentation. A white precipitate is thus formed in the limewater.
- A dilute solution of ethanol (15% alcohol) is produced. A higher alcohol content is not produced as it causes the yeast to die and fermentation to stop.
- Ethanol can be extracted by fractional distillation. This allows for higher concentrations of alcoholic beverages or pure ethanol to be produced.
- If the temperature is raised beyond 37°C, the enzymes will be denatured and fermentation will stop.

Oxidation by air

- Wine or beer turns sour when left exposed to air in the open for a few days.



- This is why fermentation of carbohydrates takes place only in the absence of oxygen.





Uses of Ethanol

- In alcoholic drinks like wine, beer or liquor.
- As a solvent in paints, varnishes, deodorants, perfumes, colognes
- As a fuel in cars (often mixed with petrol).
Ethanol is a better alternative fuel to petrol because it does not produce noxious gases. In Brazil, about 50% of the cars run on ethanol fuel, a biofuel which replaces petrol as a renewable source of energy.
- As a fuel for cooking





Carboxylic Acids

- **Carboxylic acids** are a homologous series of organic acids which have the carboxyl ($-\text{COOH}$) functional group.
- Carboxylic acids have the general molecular formula $\text{C}_n\text{H}_{2n+1}\text{COOH}$ (where $n = 0, 1, 2$ etc.). For the first member of the carboxylic acid homologous series, n in the general formula $\text{C}_n\text{H}_{2n+1}\text{COOH} = 0$. The only carbon atom in the molecule is found in the $-\text{COOH}$ functional group.
- Their names end with 'anoic acid'.

No. of C atoms	Prefix	Name (~anoic acid)	Molecular formula $\text{C}_n\text{H}_{2n+1}\text{COOH}$ ($n = \text{C atoms}$)	Full Structural formula
1	Meth-	Methanoic acid	HCOOH	$\begin{array}{c} \text{O} \\ \parallel \\ \text{H} - \text{C} \\ \backslash \\ \text{O} - \text{H} \end{array}$
2	Eth-	Ethanoic acid	CH_3COOH	$\begin{array}{c} \text{H} \\ \\ \text{H} - \text{C} - \text{C} \\ \quad \backslash \\ \text{H} \quad \text{O} = \text{O} \\ \quad \quad \backslash \\ \quad \quad \text{O} - \text{H} \end{array}$
3	Prop-	Propanoic acid	$\text{C}_2\text{H}_5\text{COOH}$	$\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H} - \text{C} - \text{C} - \text{C} \\ \quad \quad \backslash \\ \text{H} \quad \text{H} \quad \text{O} = \text{O} \\ \quad \quad \quad \backslash \\ \quad \quad \quad \text{O} - \text{H} \end{array}$
4	But-	Butanoic acid	$\text{C}_3\text{H}_7\text{COOH}$	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} \\ \quad \quad \quad \backslash \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{O} = \text{O} \\ \quad \quad \quad \quad \backslash \\ \quad \quad \quad \quad \text{O} - \text{H} \end{array}$



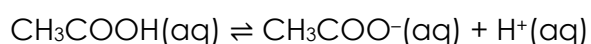
Physical Properties

- Carboxylic acids are very soluble in water.
- Their boiling points increase down the series as molecular size increases.

Chemical Properties

- Carboxylic acids are weak acids as they ionise partially in water, i.e. the concentration of H^+ ions is low (pH 5–6).

ethanoic acid \rightleftharpoons ethanoate ion + hydrogen ion



- Carboxylic acids have acidic properties due to the hydrogen ions produced.

Reactions of Carboxylic acids	Equation
Reaction with reactive metals The salts of ethanoic acid are known as ethanoates. Ethanoate ion = CH_3COO^- Salts of methanoic acid are known as methanoates, salts of propanoic acid are known as propanoates, and so on.	carboxylic acid + metal \rightarrow salt + hydrogen ethanoic acid + magnesium \rightarrow magnesium ethanoate + hydrogen $2\text{CH}_3\text{COOH}(\text{aq}) + \text{Mg}(\text{s}) \rightarrow (\text{CH}_3\text{COO})_2\text{Mg}(\text{aq}) + \text{H}_2(\text{g})$
Reaction with carbonates	carboxylic acid + carbonate \rightarrow salt + carbon dioxide + water $2\text{CH}_3\text{COOH}(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow 2\text{CH}_3\text{COONa}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$ The salt formed is sodium ethanoate.
Reaction with bases	carboxylic acid + base \rightarrow salt + water $\text{CH}_3\text{COOH}(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{CH}_3\text{COONa}(\text{aq}) + \text{H}_2\text{O}(\text{l})$ The salt formed is sodium ethanoate.
Esterification Reaction with alcohols Conc H_2SO_4 (catalyst)	carboxylic acid + alcohol \rightarrow ester + water

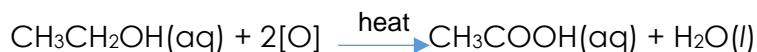


Producing Ethanoic Acid

1. Oxidation with acidified potassium manganate(VII)

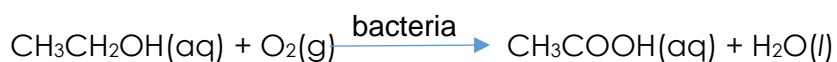
- Potassium manganate(VII) acts as an oxidising agent and is reduced. Its colour changes from purple to colourless.

Ethanol acts as a reducing agent. It is oxidised to ethanoic acid.



2. Oxidation by atmospheric oxygen

- Certain bacteria in the air oxidise ethanol to ethanoic acid.

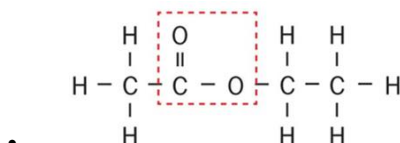


Thus, alcoholic drinks turn sour when left exposed to air for a few days.



Esters

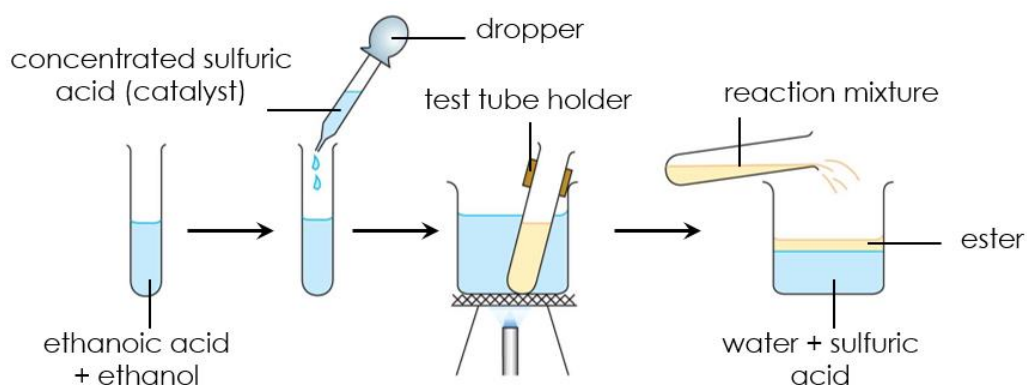
- Esters are sweet-smelling, colourless liquids that are insoluble in water.
- Esters contain the -COO- functional group.



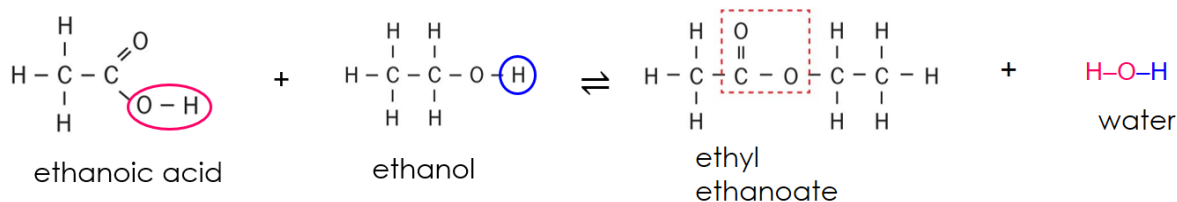
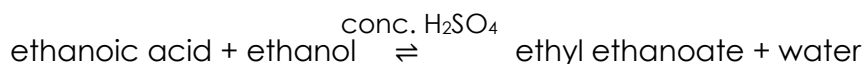
- An ester is formed by reacting a carboxylic acid with an alcohol. This reaction is known as esterification. Esterification is a condensation reaction. A condensation reaction is one where two smaller molecules join together to form a larger molecule with the elimination of a small molecule

Esterification

- An ester is formed when a mixture of an alcohol and a carboxylic acid is warmed together with a few drops of concentrated sulfuric acid.
- Concentrated sulfuric acid acts as a catalyst for the reaction.

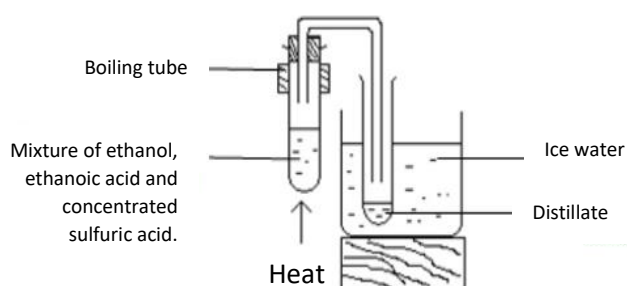
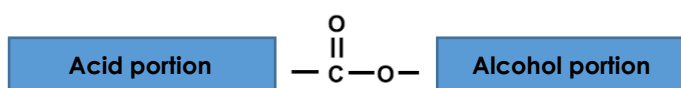


- For example, ethanoic acid and ethanol react in the presence of concentrated sulfuric acid to form water and an ester called ethyl ethanoate.





The formation of an esters can be written as



To make a sample of ethyl ethanoate in the laboratory.

1. Place a 5 cm³ ethanol (Alcohol) in a boiling tube. Then add an equal volume of ethanoic acid (Acid).
2. Swirl the boiling tube so that the liquids become well mixed.
3. Add a few drops of concentrated sulfuric acid to act as catalyst.
4. Gently heat the mixture as shown on the left.
5. The ester having a low b.p will boiled off and condensed in a separate boiling tube.

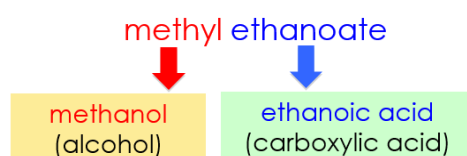


Naming an Ester

The name of an ester consists of two parts.

- The first part is derived from the alcohol.
- The second part is derived from the carboxylic acid.

Example



Name	Full Structural Formula	Organic compounds used to make esters	
		Alcohol	Organic acid
Methyl ethanoate	$ \begin{array}{c} \text{H} \quad \text{O} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{O}-\text{C}-\text{H} \\ \quad \quad \\ \text{H} \quad \quad \text{H} \end{array} $	CH ₃ OH Methanol	CH ₃ COOH Ethanoic acid
Ethyl ethanoate	$ \begin{array}{c} \text{H} \quad \text{O} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{O}-\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \\ \text{H} \quad \quad \text{H} \quad \text{H} \end{array} $	C ₂ H ₅ OH Ethanol	CH ₃ COOH Ethanoic acid
Butyl Methanoate	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{O} \\ \quad \quad \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{O}-\text{C}-\text{H} \\ \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array} $	C ₄ H ₉ OH Butanol	HCOOH Methanoic acid



FORMATION OF ESTERS

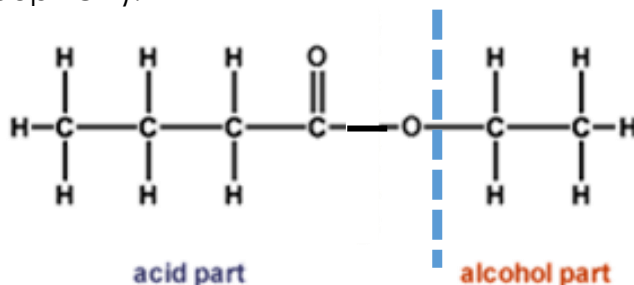
- Give the full structural formula of the ester formed between the alcohol and acid.
- Write the name of the ester, alcohol and acid below each structural formula.

Alcohol	Carboxylic acid	Esters Alcohol (~yl) Acid (~anoate)
$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{OH} \\ \\ \text{H} \end{array}$	$\begin{array}{c} \text{H} & \text{H} & \text{H} & & \text{O} \\ & & & & // \\ \text{H}-\text{C} & -\text{C} & -\text{C} & -\text{C} & \\ & & & & \backslash \\ \text{H} & \text{H} & \text{H} & & \text{O}-\text{H} \end{array}$	$\begin{array}{c} \text{H} & \text{H} & \text{H} & & \text{O} \\ & & & & // \\ \text{H}-\text{C} & -\text{C} & -\text{C} & -\text{C} & \\ & & & & \backslash \\ \text{H} & \text{H} & \text{H} & & \text{O}-\text{C}-\text{H} \\ & & & & \\ & & & & \text{H} \end{array}$
$\begin{array}{c} \text{H} & \text{H} & \text{H} \\ & & \\ \text{H}-\text{C} & -\text{C} & -\text{C}-\text{OH} \\ & & \\ \text{H} & \text{H} & \text{H} \end{array}$	$\begin{array}{c} \text{O} \\ \\ \text{H}-\text{C}-\text{OH} \end{array}$	
$\begin{array}{c} \text{H} & \text{H} \\ & \\ \text{H}-\text{C} & -\text{C}-\text{O}-\text{H} \\ & \\ \text{H} & \text{H} \end{array}$	$\begin{array}{c} \text{H} & & \text{O} \\ & & // \\ \text{H}-\text{C} & -\text{C} & \\ & & \backslash \\ \text{H} & & \text{OH} \end{array}$	
$\begin{array}{c} \text{H} & \text{H} & \text{H} & \text{H} \\ & & & \\ \text{H}-\text{C} & -\text{C} & -\text{C} & -\text{C}-\text{O}-\text{H} \\ & & & \\ \text{H} & \text{H} & \text{H} & \text{H} \end{array}$	$\begin{array}{c} \text{H} & & \text{O} \\ & & // \\ \text{H}-\text{C} & -\text{C} & \\ & & \backslash \\ \text{H} & & \text{OH} \end{array}$	
$\begin{array}{c} \text{H} & \text{H} & \text{H} \\ & & \\ \text{H}-\text{C} & -\text{C} & -\text{C}-\text{OH} \\ & & \\ \text{H} & \text{H} & \text{H} \end{array}$	$\begin{array}{c} \text{H} & & \text{O} \\ & & // \\ \text{H}-\text{C} & -\text{C} & \\ & & \backslash \\ \text{H} & & \text{OH} \end{array}$	



Deriving the alcohol and carboxylic acid from an Ester

- To name an ester we look for the part of the molecule that looks like an Acid (containing the carboxyl group - COOH).
- And the part of the molecule that looks like an Alcohol (containing the hydroxyl group -OH).



Acid part has the C = O portion from the carbonyl group. H from carboxyl group is removed.

Acid part has 4 carbon. Hence it is butanoic acid.

Alcohol part has the C - O portion from the hydroxyl group. OH from hydroxyl is removed

Alcohol part has 2 carbon. Hence it is ethanol.

Name the **alcohol** portion first then the **acid** portion.

Ester is Ethyl Butanonate



DERIVATION OF ALCOHOL AND ACID FROM ESTERS

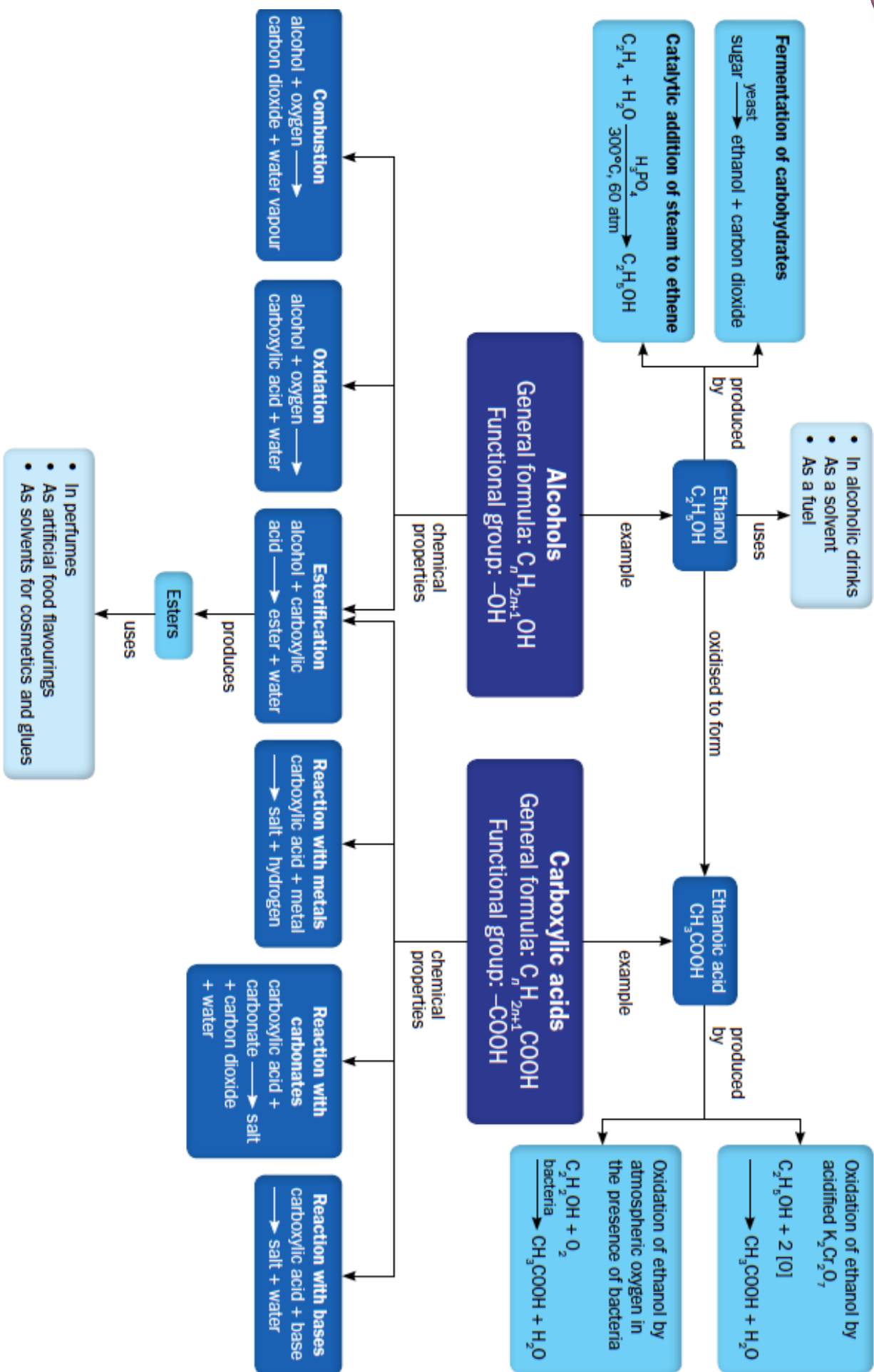
- Give the full structural formula of the the alcohol and acid.
- Write the name of the ester, alcohol and acid below each structural formula.

Esters	Alcohol	Carboxylic acid
$ \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{C} \\ \quad \quad \diagup \text{O} \\ \text{H} \quad \text{H} \quad \text{O}-\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \quad \\ \quad \quad \quad \text{H} \quad \text{H} \end{array} $		
$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C} \\ \quad \quad \quad \diagup \text{O} \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{O}-\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \quad \\ \quad \quad \quad \text{H} \quad \text{H} \end{array} $		
$ \begin{array}{c} \text{O} \\ \\ \text{H}-\text{C} \\ \diagdown \text{O} \\ \quad \text{O}-\text{C}-\text{C}-\text{H} \\ \quad \quad \\ \quad \text{H} \quad \text{H} \end{array} $		
$ \begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{C} \\ \quad \diagup \text{O} \\ \text{H} \quad \text{O}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \quad \\ \quad \quad \text{H} \quad \text{H} \quad \text{H} \end{array} $		
$ \begin{array}{c} \text{H} \quad \text{O} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{O}-\text{C}-\text{H} \\ \quad \quad \\ \text{H} \quad \quad \text{H} \end{array} $		



USES OF ESTERS

- Esters have a sweet, fruity smell. Thus, they are used in perfumes
- Used as flavourings for sweet, drinks and cakes
- Large amounts are used as solvent for glues, paints and varnishes.
- Vegetables oils and polyesters (Eg. Terylene) are also esters. Terylene (polyester) is made into fibres which are then woven into cloth - polyester clothing.
- In soaps - Naturally-occurring esters such as animal fats and vegetable oils are boiled with sodium hydroxide to form soap.





Macromolecules

A **macromolecule** is a very large molecule that is made up of many small molecules.

A **polymer** is a type of long chain macromolecule that is covalently linked together by many small repeat units called **monomers**.

Different polymers are made up of different monomers.

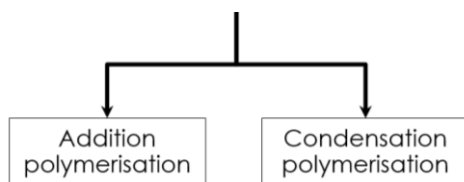
Examples of Polymers

Natural polymers	Synthetic polymers
<ul style="list-style-type: none">• proteins• DNA• starch• cellulose• wool• natural rubber• silk	<ul style="list-style-type: none">• poly(ethene)• nylon• Terylene• polyvinyl chloride (PVC)• polyvinyl acetate (PVA)• Teflon• polystyrene

Polymerisation

The process of joining together a large number of monomers to form a polymer is called **polymerisation**.

There are two types of polymerisation.





Addition Polymers

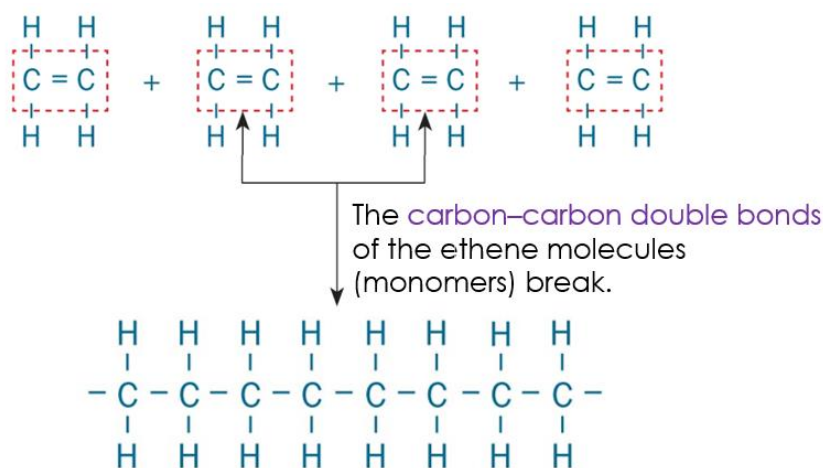
Addition polymerisation occurs when monomer units join together without losing any molecules or atoms.

Alkenes can undergo addition polymerisation.

Conditions:

- High pressure (e.g. 1000 atm)
- High temperature (e.g. 200°C)
- Catalyst

Addition Polymer: Poly(ethene)



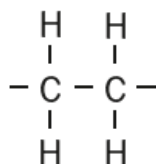
Each monomer forms **single bonds** with two other monomers. Eventually, they join to form the **polymer**, poly(ethene).

3 Equation for Polymerisation of Ethene

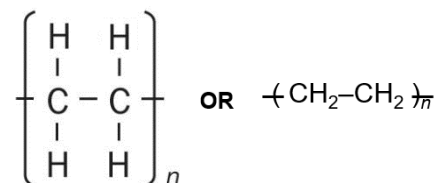
1	Word equation	ethene $\xrightarrow[\text{catalyst}]{\text{high temperature and pressure}}$ poly(ethene)
2	Structural formula	$n (\text{CH}_2=\text{CH}_2) \xrightarrow[\text{catalyst}]{\text{high temperature and pressure}} \text{+CH}_2\text{-CH}_2\text{+}_n$
3	Full structural formula	$n \begin{array}{c} \text{H} & \text{H} \\ & \\ \text{C} & = & \text{C} \\ & \\ \text{H} & \text{H} \end{array} \xrightarrow[\text{catalyst}]{\text{high temperature and pressure}} \left[\begin{array}{cc} \text{H} & \text{H} \\ & \\ \text{C} & - & \text{C} \\ & \\ \text{H} & \text{H} \end{array} \right]_n$

**Repeat Unit of Poly(ethene)**

- The repeat unit of poly(ethene) is:



- The formula of the polymer can be written using its repeat unit. Thus, the formula of poly(ethene) can be written simply as:



- The letter 'n' represents a large number.
- The number of ethene molecules in each poly(ethene) molecule varies between 10 000 and 30 000.

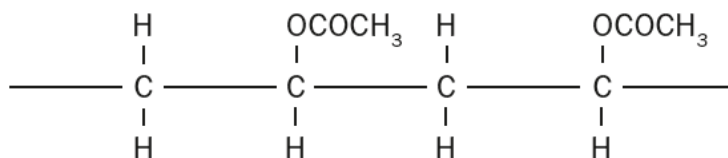


Deducing the Structural Formula of a Monomer from Its Polymer

Example

PVA (polyvinyl acetate) is used as the main constituent in emulsion paints.

It has the structural formula:



Step	Explanation
1	<p>Identify the repeat unit in the polymer.</p> $\begin{array}{ccccccc} & \text{H} & & \text{OCOCH}_3 & & \text{H} & & \text{OCOCH}_3 \\ & & & & & & & \\ \text{---} & \text{C} & \text{---} & \text{C} & \text{---} & \text{C} & \text{---} & \text{C} & \text{---} \\ & & & & & & & \\ & \text{H} & & \text{H} & & \text{H} & & \text{H} \end{array}$ <p style="text-align: center;"> ← repeat unit → ← repeat unit → </p>
2	<p>Write down the formula of the repeat unit.</p> $\begin{array}{ccccccc} & \text{H} & & \text{OCOCH}_3 & & \text{H} & & \text{OCOCH}_3 \\ & & & & & & & \\ \text{---} & \text{C} & \text{---} & \text{C} & \text{---} & \text{C} & \text{---} & \text{C} & \text{---} \\ & & & & & & & \\ & \text{H} & & \text{H} & & \text{H} & & \text{H} \end{array}$ <p style="text-align: center;"> ← repeat unit → ← repeat unit → </p> <p style="text-align: center;">↓</p> $\begin{array}{ccccc} & \text{H} & & \text{OCOCH}_3 \\ & & & \\ \text{---} & \text{C} & \text{---} & \text{C} & \text{---} \\ & & & \\ & \text{H} & & \text{H} \end{array}$
3	<p>Convert the carbon-carbon single bond into a carbon-carbon double bond.</p> <p>The structural formula of the monomer, vinyl acetate, is</p> $\begin{array}{ccc} \begin{array}{ccccc} & \text{H} & & \text{OCOCH}_3 \\ & & & \\ \text{---} & \text{C} & \text{---} & \text{C} & \text{---} \\ & & & \\ & \text{H} & & \text{H} \end{array} & \longrightarrow & \begin{array}{ccccc} & \text{H} & & \text{COOCH}_3 \\ & & & \\ & \text{C} = & \text{C} & \\ & & & \\ & \text{H} & & \text{CH}_3 \end{array} \end{array}$
4	<p>To draw the polymer, simply use the repeating unit and put a bracket and 'n'</p> $\left[\begin{array}{ccccc} & \text{H} & & \text{OCOCH}_3 \\ & & & \\ \text{---} & \text{C} & \text{---} & \text{C} & \text{---} \\ & & & \\ & \text{H} & & \text{H} \end{array} \right]_n$



Uses of Addition Polymers

Polymer	Properties	Uses
Poly(ethene)	<ul style="list-style-type: none"> Poly(ethene) can be easily moulded into different shapes. 	<ul style="list-style-type: none"> Poly(ethene) is used to make <ul style="list-style-type: none"> plastic toys; clingfilm; plastic bags; buckets.
Polystyrene	<ul style="list-style-type: none"> Polystyrene is hard, light and brittle 	<ul style="list-style-type: none"> It is used to make disposable containers.
Polyvinyl chloride (PVC)	<ul style="list-style-type: none"> Resistant to chemical attack 	<ul style="list-style-type: none"> Polyvinyl chloride (PVC) is used to make pipes, raincoats, flooring mats and thin gloves.
Teflon (polytetrafluoroethene)	<ul style="list-style-type: none"> Non stick 	<ul style="list-style-type: none"> Teflon (polytetrafluoroethene) is used in 'non-stick' frying pans.

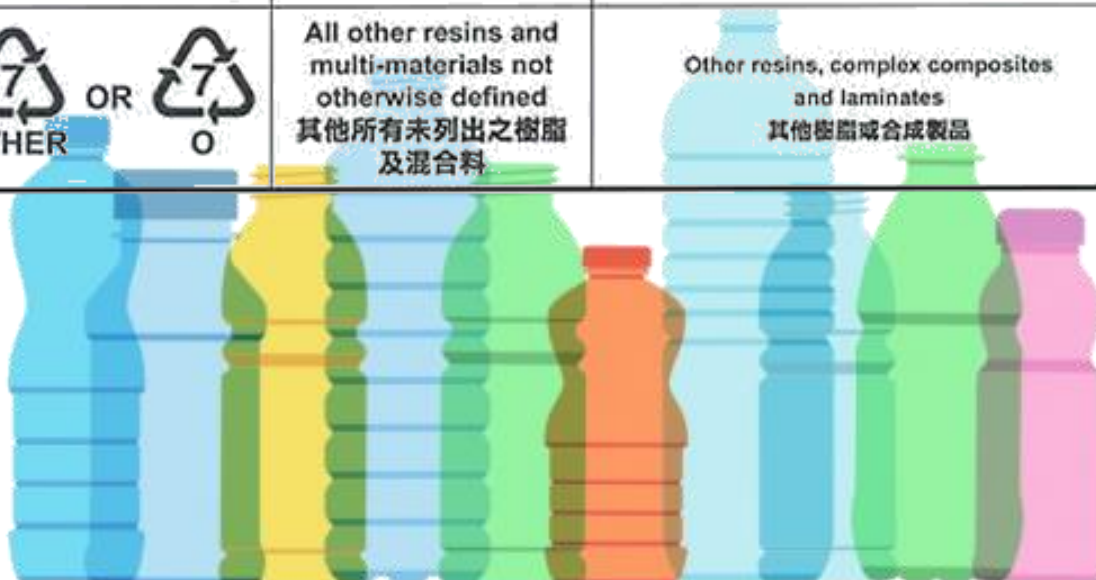
Addition polymers are also good insulators of heat and electricity,





The Plastic Coding System

Code 編碼	Material 物料	Applications 應用例子
 OR  PET OR PETE	Polyethylene Terephthalate 聚對苯二甲酸乙二醇酯	Clear soft drink and beverage bottles, food packaging 透明汽水及飲品樽，食品包裝
 OR  HDPE OR PE-HD	High Density Polyethylene 高密度聚乙烯 (硬性軟膠)	Bottles (especially for food products, detergent and cosmetics), industrial wrapping and film, sheets, plastic bags 食物，洗潔精及化妝品樽，工業包裝及薄膜，背心膠袋
 OR  PVC OR V	Polyvinyl Chloride 聚氯乙烯	Bottles, packaging film, credit cards, water containers, water pipes 塑膠樽，包裝薄膜，信用卡，盛水容器，水管
 OR  LDPE OR PE-LD	Low Density Polyethylene 低密度聚乙烯	Cling film, plastic bags, flexible containers and food wrap 保鮮膜，背心膠袋，彈性容器，食品包裝
 PP	Polypropylene 聚丙烯 (百折膠)	Packaging such as yoghurt and margarine pots, sweet and snack wrappers, medical packaging, milk and beer crates, shampoo bottles 酸乳酪及牛油器皿，糖果及小吃包裝，醫療用品包裝，牛奶及啤酒樽箱，洗頭水樽
 PS	Polystyrene 聚苯乙烯 (硬膠)	Disposable hot or cold drink cups and plates, fast food clamshells, dairy product containers 塑膠杯碟，外賣飯盒，乳製品容器
 OR  OTHER OR O	All other resins and multi-materials not otherwise defined 其他所有未列出之樹脂及混合料	Other resins, complex composites and laminates 其他樹脂或合成製品





Condensation Polymers

Some polymers are made by reacting two different types of monomers.

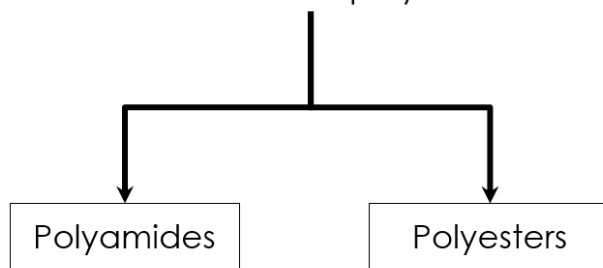
Each of the monomers involved has a functional group at each end of the molecule.

When these monomers react, a polymer is produced.

A small molecule, such as water, is also produced as a by-product of the reaction.

Condensation polymerisation occurs when monomers combine to form a polymer, with the removal of a small molecule such as water.

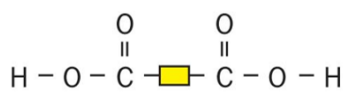
There are two main groups of condensation polymers.



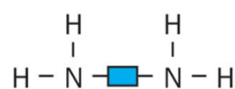
Nylon — A Synthetic Polyamide

Nylon is the first synthetic fibre made by condensation polymerisation.

Its monomers are a dicarboxylic acid and a diamine.



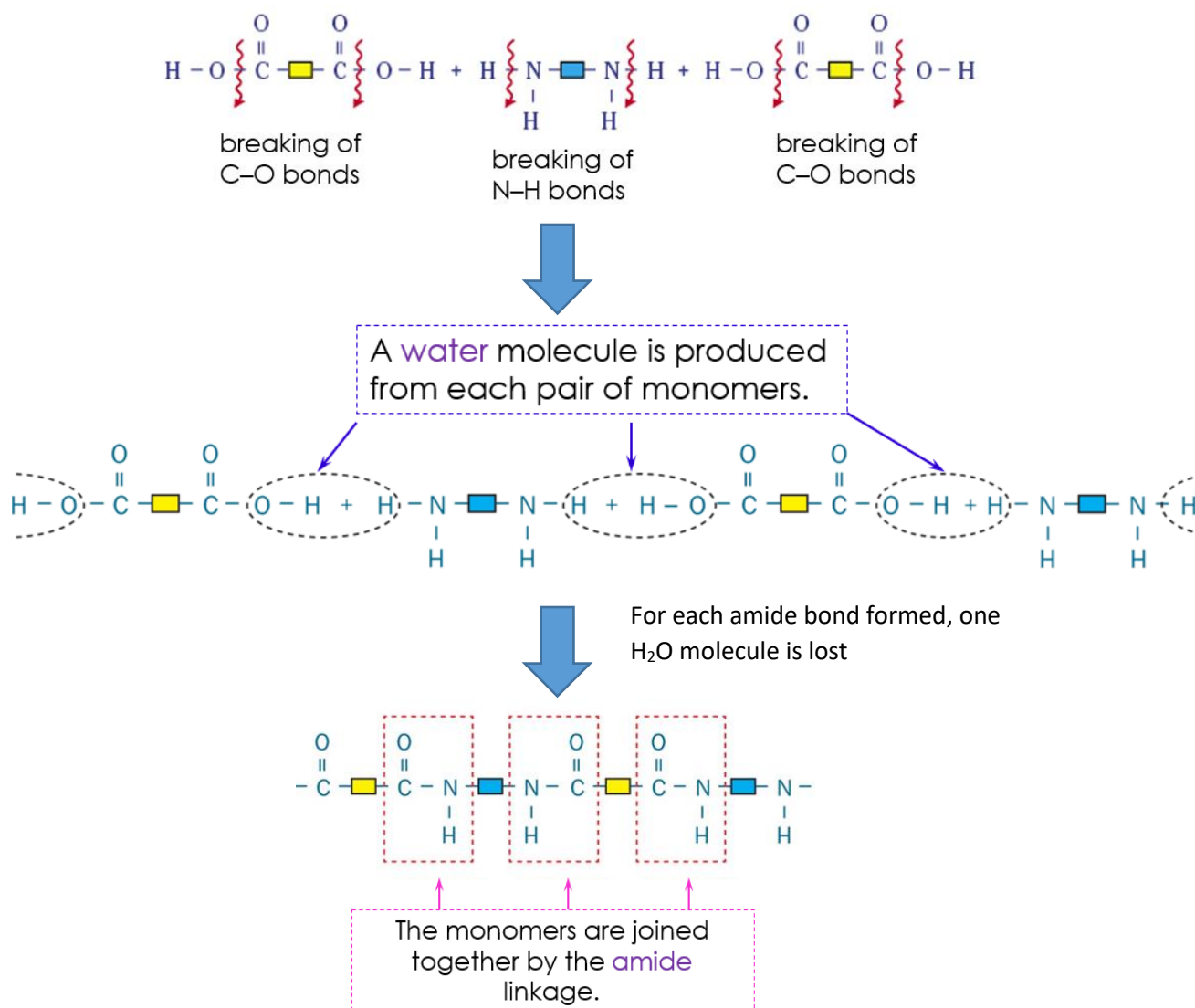
Dicarboxylic acid



Diamine

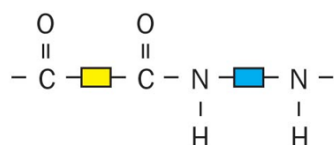


This is how the two monomers react:



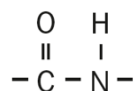
Repeat Unit and Functional Group of Nylon

The repeat unit of nylon is:



Nylon is a polyamide as it contains many amide linkages.

The amide linkage is usually written as:

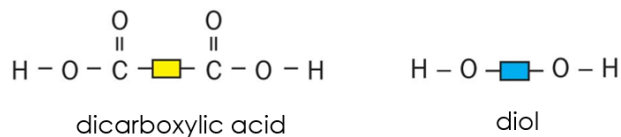




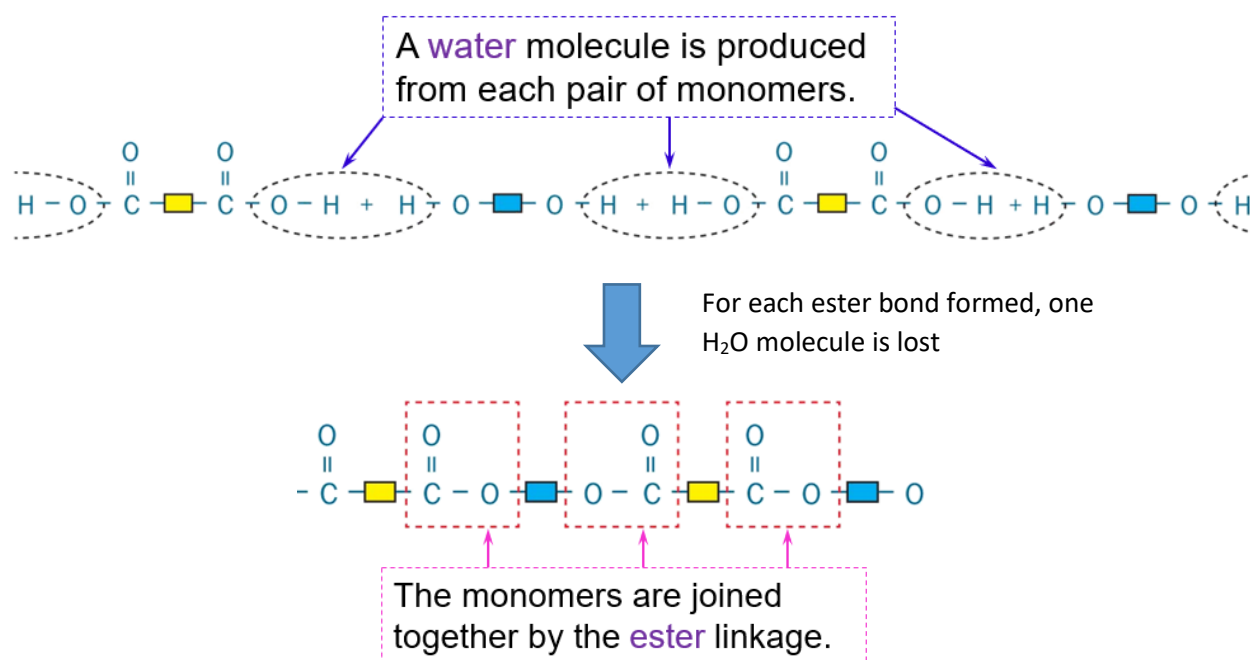
Terylene — A Synthetic Polyester

Terylene is another example of a condensation polymer.

Terylene is made from these monomers:

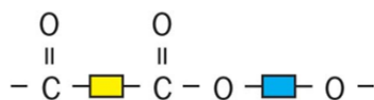


This is how the two monomers react:



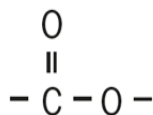
Repeat Unit and Functional Group of Terylene

The repeat unit of Terylene is:



Terylene is a polyester as it contains many ester linkages.

The ester linkage is usually written as:





Uses of Man-made Fibres

- Synthetic or man-made fibres are usually plastics which have been spun into threads.
- Terylene and nylon are the best known examples of synthetic fibres.
- Clothes made from these fibres are shrink-proof and crease-proof. They are also easier to wash and dry.
- Examples of items made from nylon and Terylene are curtains, parachutes, fishing lines and sleeping bags.





Plastics and Pollution

Advantages of Using Plastics

Plastics are increasingly being used in place of natural materials such as wood, metal, cotton and leather.

Reasons:

- Relatively cheap
- Easily moulded into various shapes
- Light, tough and waterproof
- Durable



Pollution Problems Caused by Plastics

Land Pollution	<ul style="list-style-type: none"> • Plastics are non-biodegradable and are not decomposed by bacteria in the soil. • This leads to a build up of waste.
Air Pollution	<ul style="list-style-type: none"> • Plastics are flammable. • When plastics burn, fires can spread very quickly and poisonous gases are produced.
Water pollution	<ul style="list-style-type: none"> • Plastics thrown into the sea endanger marine animals. These animals mistaken plastics for food and choke on them. • Plastics can clog up rivers and drains, which may become breeding grounds for mosquitoes.

How Can We Properly Dispose of Plastics?

Many plastics produce poisonous gases upon incineration.

For example, PVC produces hydrogen chloride gas on burning.

Burning plastics will therefore cause air pollution.

The best way to deal with plastic objects is to reuse or recycle them.



