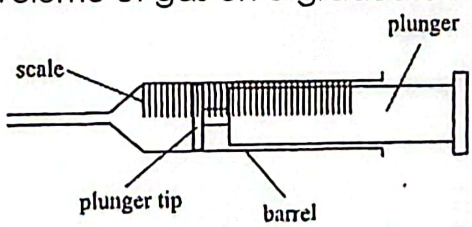
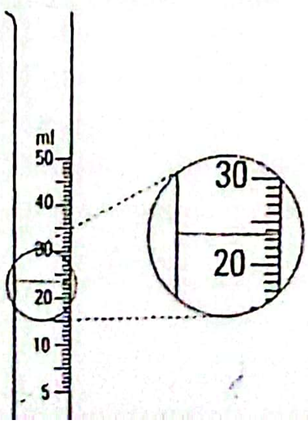
Experimental Design

MEASUREMENT

Physical properties	SI Units		Other common units	Common Measuring Apparatus
	Name	Symbol		
Time	second	s	• hour (h) • minute (min) 1 h = 60 min = 3600 s	• digital stopwatch
Temperature	kelvin	K	• degree Celsius (°C)	• mercury-in-glass thermometer • alcohol-in-glass thermometer
Mass	kilogram	kg	• gram (g) • milligram (mg) • tonne (t) 1 kg = 1000 g = 1000000 mg	• electronic balance
Volume	cubic metre	m^3	• cubic centimetre (cm^3) • cubic decimetre (dm^3) 1 dm^3 = 1000 cm^3	Volume of **liquids**: • burette* • pipette • measuring cylinder Volume of **gases**: • gas syringe *(See below for more information)*

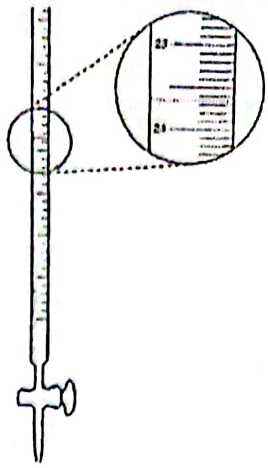
MEASURING VOLUME

Measuring Volume of	
Gas	**(1) Gas Syringe** • Used to collect gases of any density. • It measures the volume of gas on a graduated scale. 
Liquids or Solutions	**(1) Measuring Cylinder** • Measures approximately volumes of a liquid to the nearest 0.5 cm^3. Eg: 21.0 cm^3 or 15.5 cm^3 

(2) Burette

- Has an accuracy of 0.1 cm^3.
- Accurately measures the volume of a liquid to the nearest 0.05 cm^3.

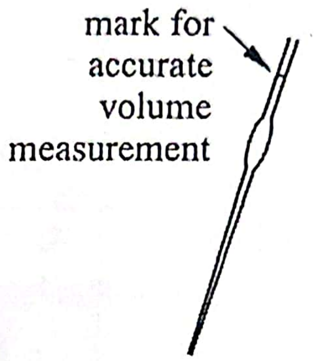
Eg: 23.65 cm^3



(3) Pipette

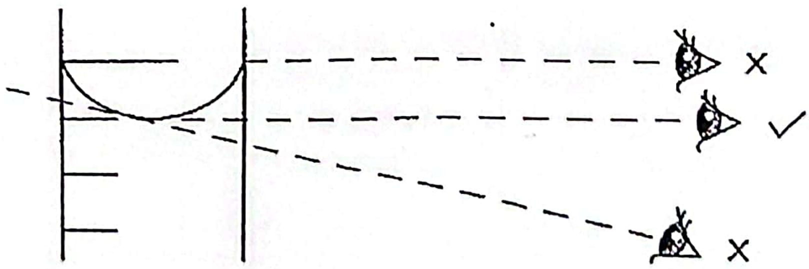
- Accurately measures a <u>fixed volume</u> of liquid.

Eg: 25.0 cm^3

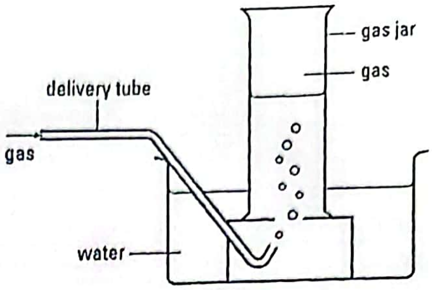
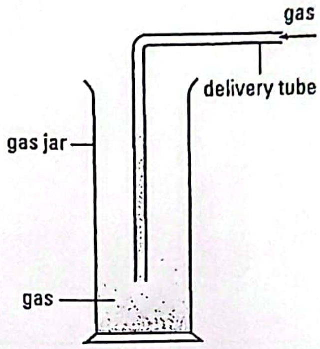
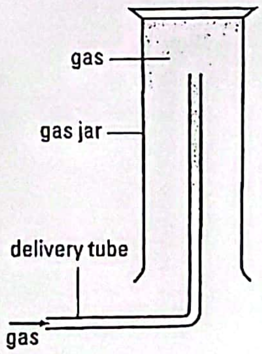
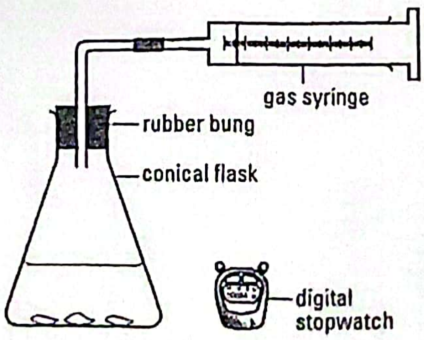


MENISCUS

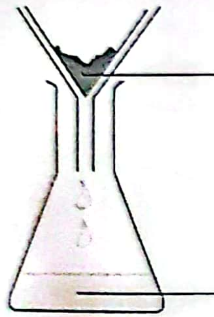
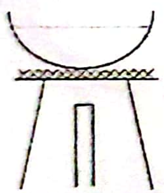
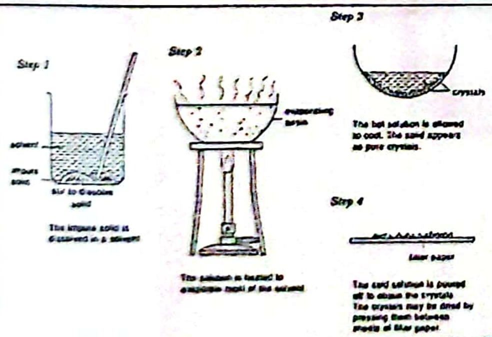
- The surface of the liquid that curves at the side of the measuring instrument.
- To get an accurate volume reading, it must be read from the **lowest level or the highest level of the meniscus.**

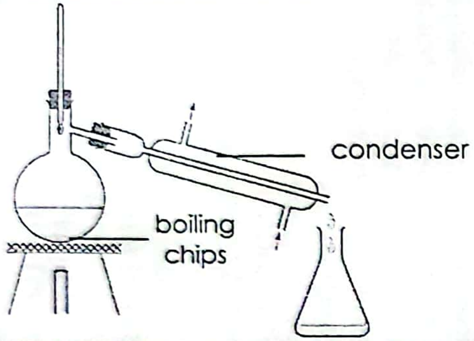
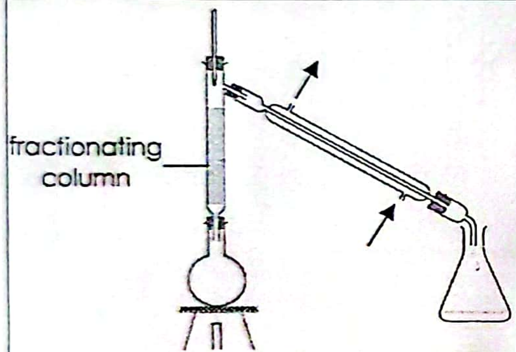
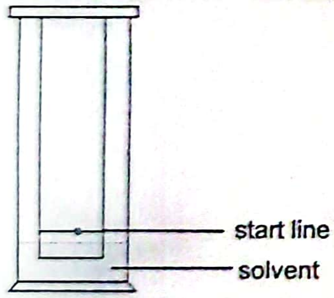


<u>GAS COLLECTION</u>

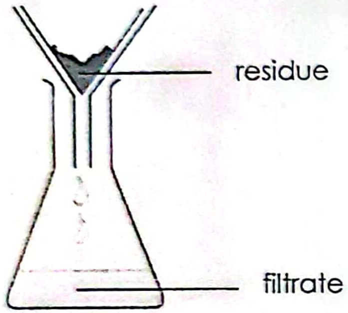
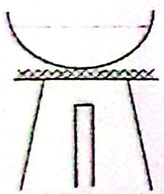
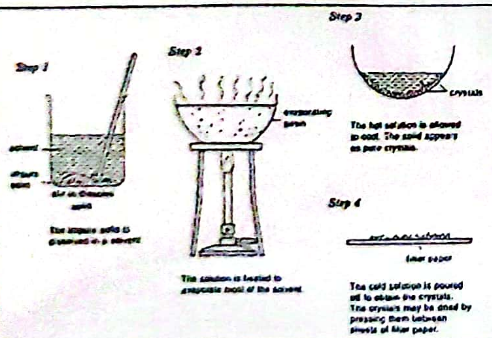
Methods		
(1) Displacement of Water	Used to collect gases that are <u>insoluble or slightly soluble in water</u>, and have a density almost the same as air. Eg: hydrogen (H_2), oxygen (O_2), carbon dioxide (CO_2).	
(2) Displacement of Air	Used to collect gases that are <u>soluble in water</u>. • **Downward Delivery** Used for soluble gases that are <u>denser than air</u>. Eg: chlorine (Cl_2), hydrogen chloride (HCl), sulfur dioxide (SO_2). • **Upward Delivery** Used for soluble gases that are <u>less dense than air</u>. Eg: hydrogen (H_2), ammonia (NH_3).	 
(3) Using a Gas Syringe	Gas produced can be collected in a gas syringe to measure the volume of gas.	 • E.g.: collecting the volume of gas produced at fixed time interval to determine the speed of a reaction

Separation Technique for <u>Mixtures</u>

Technique	Diagram	Mixture it Separates	Basis of Separation	Applications
Filtration		• Insoluble solids from liquids • Insoluble solids from soluble solids after a suitable solvent is added	• Pores on filter paper do not allow big particles to pass through	• Mud/ sand from water • Chalk + sugar/ salt/ copper(II) sulfate (after dissolving the soluble substances in water) • *Salts prep*: removing excess insoluble substance from the acid is fully reacted • *Salts prep*: <u>crystals</u> after crystallisation is complete • *Salts prep*: <u>precipitate</u> as the residue after mixing 2 solutions
Evaporation (to dryness)		• Soluble solids from liquids that are **not decomposed by heat**	• Solvent/ liquid has a lower boiling point from solids that allows the solvent to be removed	• Obtain <u>table salt (sodium chloride)</u> from sea water
Crystallisation		• Soluble solids from liquids that are **decomposed by heat** (e.g. sugar, copper(II) sulfate)	• Different solubility of substances at different temperature of solvent/ liquid (e.g. hot water can dissolve a higher mass of solute than colder water)	• Obtaining <u>sugar</u> from sugar solution • *Salts prep*: Obtaining soluble salts as crystals from a hot, saturated salt solution e.g. <u>copper(II) sulfate crystals</u> from copper(II) sulfate solution

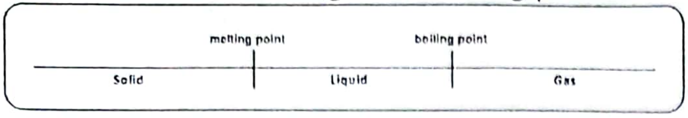
Simple distillation		• Solids/solutes from solvents	• Different boiling points	• <u>Water</u> (as distillate) from sea water
Fractional distillation		• Miscible liquids with different boiling points	• Different boiling points	• Fractional distillation of - Liquefied air - Petroleum/crude oil - During fermentation: to obtain <u>ethanol solution</u> from a mixture of glucose, water, yeast
Paper chromatography		• Substances with **different solubility in the same solvent** (A pure substance produces only one spot on the chromatogram.)	• Different solubility	• Drug tests • Analysis of - food colouring - artificial food flavouring - dyes in inks

Separation Technique for Mixtures

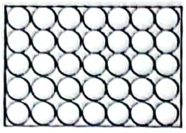
Technique	Diagram	Mixture it Separates	Basis of Separation	Applications
Filtration		• Insoluble solids from liquids • Insoluble solids from soluble solids after a suitable solvent is added	• Pores on filter paper do not allow big particles to pass through	• Mud/ sand from water • Chalk + sugar/ salt/ copper(II) sulfate (after dissolving the soluble substances in water) • *Salts prep*: removing excess insoluble substance from the acid is fully reacted • *Salts prep*: crystals after crystallisation is complete • *Salts prep*: precipitate as the residue after mixing 2 solutions
Evaporation (to dryness)		• Soluble solids from liquids that are **not decomposed by heat**	• Solvent/ liquid has a lower boiling point from solids that allows the solvent to be removed	• Obtain table salt (sodium chloride) from sea water
Crystallisation		• Soluble solids from liquids that are **decomposed by heat** (e.g. sugar, copper(II) sulfate)	• Different solubility of substances at different temperature of solvent/ liquid (e.g. hot water can dissolve a higher mass of solute than colder water)	• Obtaining sugar from sugar solution • *Salts prep*: Obtaining soluble salts as crystals from a hot, saturated salt solution e.g. copper(II) sulfate crystals from copper(II) sulfate solution

Kinetic Particle Theory:

Physical state at room temperature (25 °C) depends on the boiling and melting points.

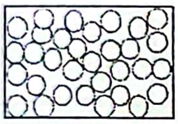


Solid:



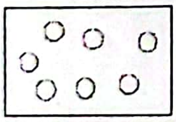
- Particles very closely packed together in an orderly arrangement
- Particles are vibrating about their fixed positions
- Very strong attractive forces between particles
- Fixed volume and fixed shape, cannot be compressed

Liquid:

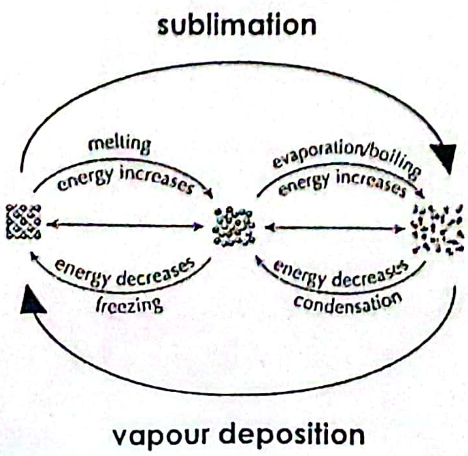


- Particles closely packed together but not in an orderly arrangement
- Particles slide over one another
- Strong attractive forces between particles
- Fixed volume but no fixed shape, cannot be compressed

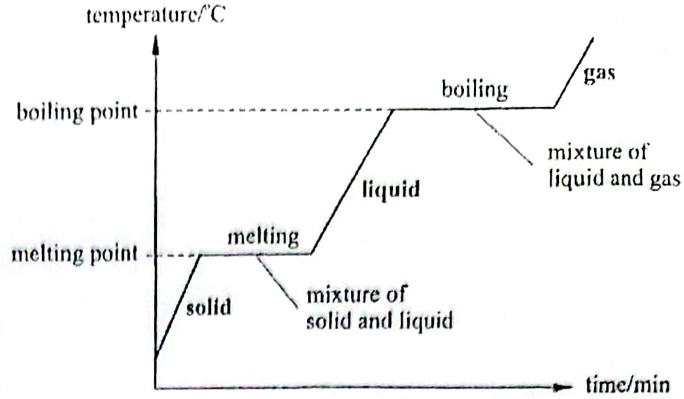
Gas:



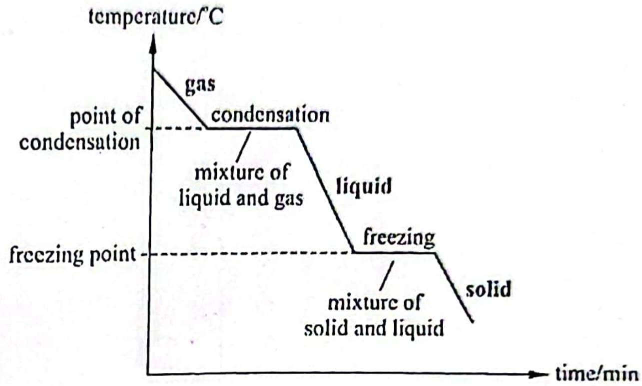
- Particles are spread very far apart, in random arrangement
- Particles moving about freely at high speeds.
- Very weak attractive forces between particles
- No fixed volume and no fixed shape, can be compressed easily



Heating curve:



Cooling curve:



Element:

A pure substance that cannot be broken down into two or more simpler substances by chemical methods.

- **Fixed melting & boiling points**
  e.g. copper, iron, hydrogen, oxygen

Compound:

A pure substance made up of two or more different elements chemically combined together.

- **Fixed melting & boiling points**
- Can only be separated by chemical methods / chemical reactions
- Fixed proportion by mass
- Properties different from elements that make up the compound.
  e.g. calcium chloride, sulfur dioxide

Mixture:

An impure substance made up of two or more substances that are not chemically combined together.

- No fixed melting & boiling points, **melts and boils over a range of temperatures**
- Can be separated by physical methods (separation techniques)
- Variable proportion by mass
- Properties same as substances that make up the mixture.
  e.g. air, alloys like steel & brass, petroleum

Nucleon / mass number:
number of neutrons + number of protons

Proton / atomic number:
number of protons

An atom:

		Relative charge	Relative mass
Nucleus	Neutron	0	1
	Proton	1+	1
Shells	Electron	1–	1/1840

Isotopes:
Atoms of the same element, having **the same number of protons** (and electrons) but **different number of neutrons.**

- **similar chemical reactions** (as the isotopes have the same number of valence electrons) but
- **different physical properties** such as m.p. and b.p. due to different masses (the isotopes have different number of neutrons).

Explain why the (A_r) relative atomic mass of chlorine is not a whole number.

- Chlorine exist as two different isotopes *Cl*–35 (75%) and *Cl*–37 (25%)
- The relative atomic mass is the average weighted mass of all the isotopes of chlorine (35 x 75/100) + (37 x 25/100) = 35.5

Metallic atoms

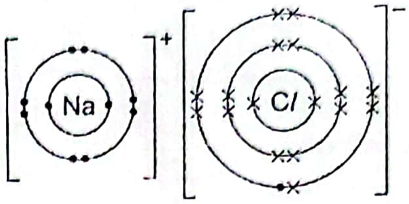
- Tend to have 1, 2 or 3 valence electron(s).
- Metallic atoms will lose their valence electron(s) to obtain a fully-filled valence electron shell like noble gases to form a positively-charged ion (protons > electron)

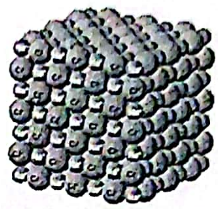
Non-metallic atoms

- Tend to have 5, 6 or 7 valence electron(s).
- Non-metallic atoms will gain electron(s) from metals to obtain a fully-filled valence electron shell like noble gases to form a negatively-charged ion (electrons > protons)

Dot-and-cross diagrams of:

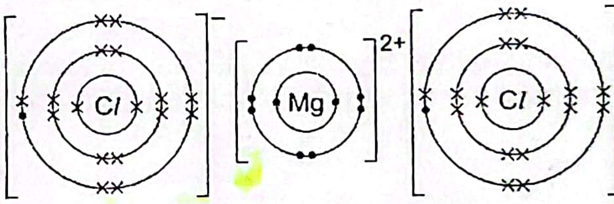
- Sodium chloride, NaC*l*





Giant ionic crystal lattice structure

- Magnesium chloride, MgC*l*$_2$



Ionic Bonding:

- Strong electrostatic forces of attraction between oppositely-charged ions.
- Involves the **transfer of valence electron(s) from metals to non-metals.**

Physical Properties of Ionic Compounds

1. **High melting & boiling points (large amount of energy required to overcome strong electrostatic forces of attraction between oppositely-charged ions)** – usually solids at r.t.p.
2. **Able to conduct electricity in molten & aqueous states (due to free mobile ions that carry charges), but not in the solid state as ions are held in fixed positions by strong electrostatic forces of attraction**
3. Usually soluble in water, insoluble in organic solvents

Ions:

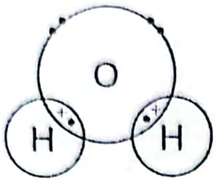
hydroxide ion: OH^-	ammonium: ion: NH_4^+
nitrate ion: NO_3^-	silver ion: Ag^+
sulfate ion: SO_4^{2-}	zinc ion: Zn^{2+}
carbonate ion: CO_3^{2-}	

Covalent Bonding:

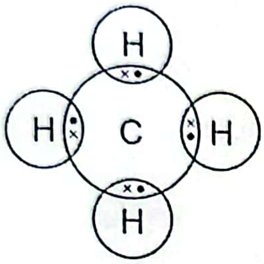
- Involves the **sharing** of electron(s) **between non-metallic atoms.**

Dot-and-cross diagrams of:

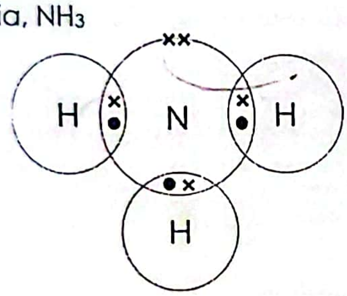
- Water, H_2O



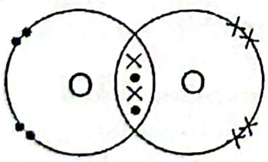
- Methane, CH_4



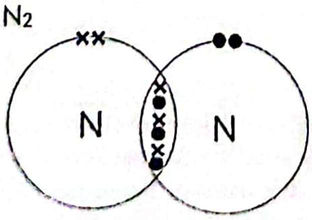
- Ammonia, NH_3



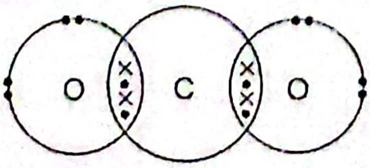
- Oxygen, O_2



- Nitrogen, N_2



- Carbon dioxide, CO_2



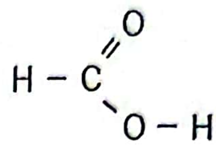
Q. If shown the **structural formula** of carbon dioxide, what is the number of electrons shared?

$$O=C=O$$

Workings:

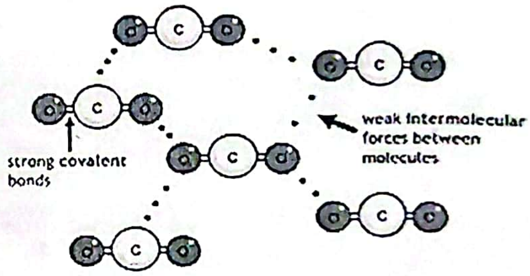
4 covalent bonds x 2 e^- = **8 electrons shared**

Q. What is the number of electrons shared in a molecule of methanoic acid?



Physical Properties of simple covalent molecules:

E.g.



1. Low melting & boiling points (**small amount of energy required to overcome the weak intermolecular forces of attraction between molecules)** – usually **liquids** or **gases** at r.t.p.
2. Does not conduct electricity in any state **(no free mobile ions AND electrons)**
   Exceptions: aqueous acids, NH_3 (aq)
3. Insoluble in water, soluble in organic solvents (except, acids, NH_3, alcohol, glucose)

Relative atomic mass, A_r:

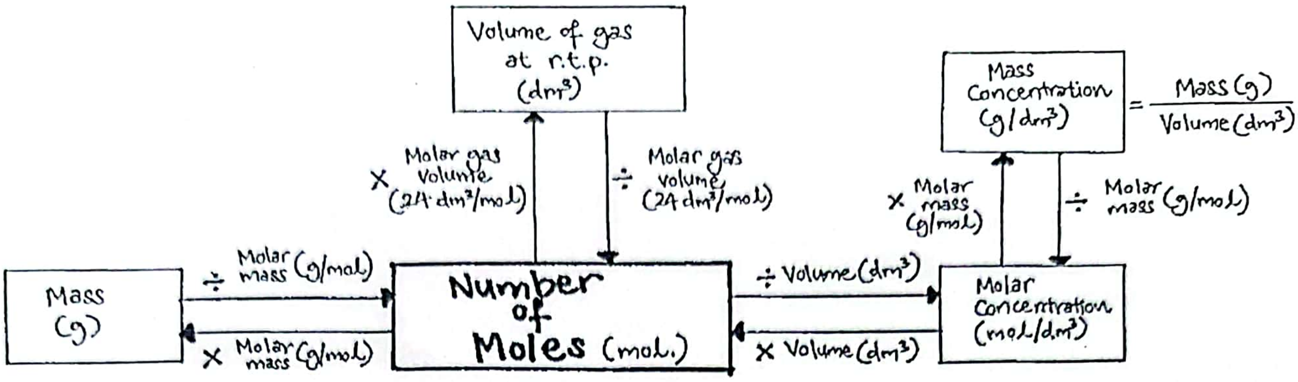
- The relative atomic mass of an element is the average mass of all the isotopic atoms present in the element compared to 1/12 the mass of a carbon-12 atom.

Relative molecular mass, M_r:

- The relative molecular mass of a molecule is the average mass of the molecule of an element or compound compared to 1/12 the mass of a carbon-12 atom.

- **M_r = sum of the A_r of all the atoms in the molecule**

Mole Concept:



Conversion: $1\ dm^3 = 1000\ cm^3$ 1 kg = 1 000 g   1 tonne = 1 000 000 g

Limiting reactant/reagent:
- The reactant that is completely used up in the reaction; the other reactant is in excess
- **Determines the no. of mol. of product(s) formed in a reaction.**

Acid:
- A substance that dissociates in water to produce hydrogen ions **(H^+)**.
  - **Strong** acids (Dissociate **completely** in water to produce a lot of H^+ ions), eg:
    $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
    $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$
    $H_2SO_4 \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$
  - Weak acids (Dissociate **partially** in water to produce very little H^+ ions), eg: CH_3COOH

Physical Properties of Acids:
- tastes sour
- pH < 7
- turns blue litmus paper red
- conducts electricity in aqueous solution (due to presence of mobile ions in solution)

Base (Metal Oxide / Metal Hydroxide):
- Bases that are soluble in water are also known as alkalis

Alkali:
- A substance that dissociates in water to produce hydroxide ions **(OH^-)**.
  - Strong alkali, eg:
    $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$
    $KOH(aq) \rightarrow K^+(aq) + OH^-(aq)$
  - Weak alkali, eg: aqueous NH_3

Physical Properties of Alkalis:
- tastes bitter
- feels soapy & slippery
- pH > 7
- turns red litmus paper blue
- conducts electricity in aqueous solution (due to presence of mobile ions in solution)

Chemical Reactions:
1. **Acid** + **Alkali** → Salt + Water ***(Neutralisation)***
   Ionic equation: $\mathbf{H^+(aq) + OH^-(aq) \rightarrow H_2O(l)}$
2. **Acid** + Base → Salt + Water
3. **Acid** + Carbonate
   → Salt + Water + Carbon dioxide
4. **Acid** + Metal (except copper, silver, gold)
   → Salt + Hydrogen gas
5. **Alkali** + Ammonium salt
   → Salt + Water + Ammonia

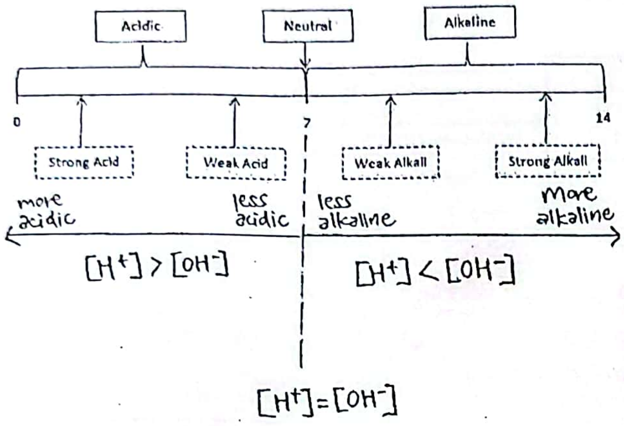
Test for **hydrogen (H_2):**
- Test: Place a **lighted splint** at the mouth of the test tube.
- Observation: Lighted splint **extinguished with a 'pop' sound.**

Test for **carbon dioxide (CO_2):**
- Reagent: Limewater, $Ca(OH)_2$
- Test: **Bubble the gas into limewater.**
- Observation: **White precipitate** (of $CaCO_3$) is produced in limewater.

Test for **ammonia (NH_3):**
- Test: Place **moist red litmus paper** near the mouth of the test tube.
- Observation: Moist red litmus paper turns blue if ammonia is present.
  (Note: ammonia is a pungent gas.)

pH Scale:

- 0 to 14
- The lower the pH, the more acidic the solution, the higher the concentration of H^+ ions.
- The higher the pH, the more alkaline the solution, the higher the concentration of OH^- ions.



Common indicators (unable to show pH of solution):

indicator	colour in acidic solution	colour in alkaline solution
Methyl orange	red	yellow
Litmus	red	blue

Universal Indicator (U.I.):
pH < 7, Acidic: Red, Orange, Yellow
pH = 7, Neutral: Green
pH > 7, Alkaline: Blue, Indigo, Violet

- **HCl, H_2SO_4, HNO_3 – pH 0 to 2 U.I. colour red**
- **Aqueous NH_3 – pH 10, U.I. colour blue**
- **NaOH, KOH – pH 13 or 14 U.I. colour violet**

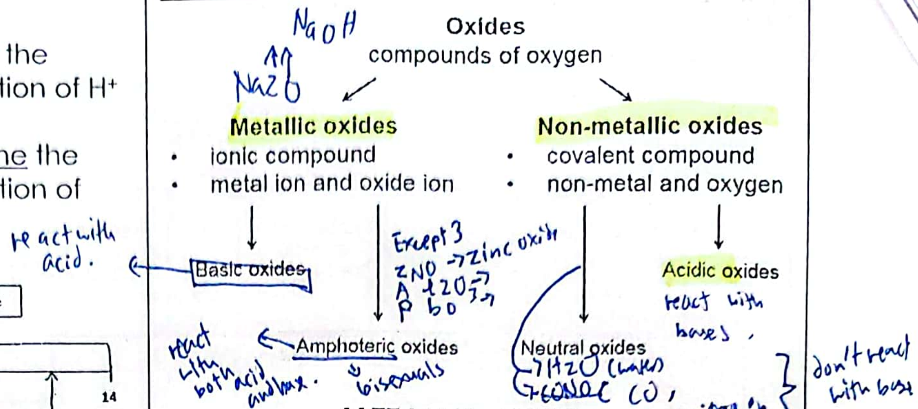
(Note: acid rain has an approx. pH of 4.)

pH meter/ pH sensor connected to datalogger: more accurate and reliable.

What can be added to acidic soil to reduce acidity (neutralise the excess acids in the soil)?

- **Calcium hydroxide** (slaked lime)

Oxides:



METALLIC OXIDES

Basic oxides:

- React with acids to form salt and water.
- Eg: CaO, Na_2O

Amphoteric oxides:

- React with acids AND bases to form salt and water.
- Eg: **ZnO, Al_2O_3 , PbO (ZAP)**

NON-METALLIC OXIDES

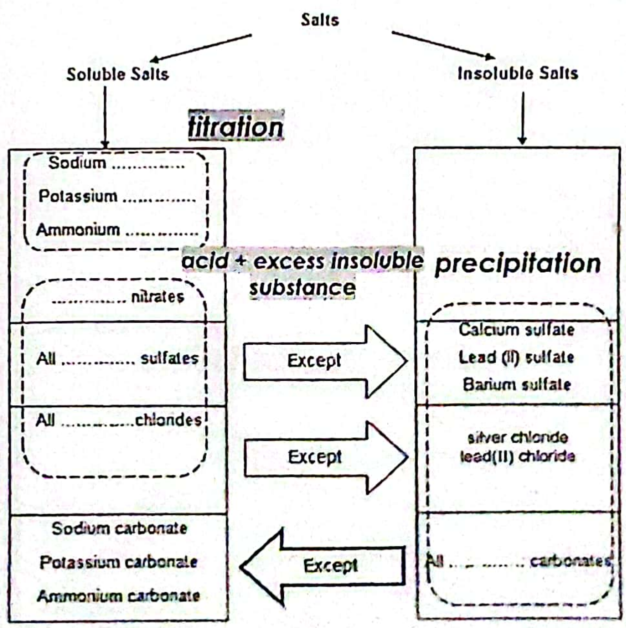
Acidic oxides:

- React with bases/alkalis to form salt and water.
- Eg: CO_2, SO_2. NO_2

Neutral oxides:

- No reaction with acids and bases.
- Eg: **NO, CO, H_2O** (1 O atom only)

Solubility Table of Salts: ★



g foss'

How are insoluble salts prepared?

- **By precipitation**
  soluble + soluble → insoluble salt
  - Two solutions: one containing the cation and one containing the anion are mixed;
  - the precipitate is obtained by filtration as the residue; the excess reagent solution will be in the filtrate.

E.g.
Barium sulfate, $BaSO_4$
$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$

1. Add excess sodium sulfate solution into barium nitrate solution while stirring. A precipitate of barium sulfate is formed.
2. Filter the mixture to collect the precipitate as the residue.
3. Wash the precipitate with a little cold distilled water to remove the impurities.
4. Allow the precipitate to dry on a piece of filter paper.

How are most soluble salts prepared?

- **Reactions of acids with insoluble substances**
  insoluble + soluble (acid) → soluble salt
- The excess reactant ***ensures that the acid is completely reacted***. Once the reaction is complete, the excess solid can be filtered off to obtain a solution of the soluble salt.
  - acid + excess reactive metal (**✓ Mg, Al, Zn, Fe**)
  - acid + excess insoluble carbonate
  - acid + excess insoluble base

E.g.
acid + excess reactive metal
$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

1. Add excess zinc to dilute sulfuric acid, until no more zinc dissolves.
2. Filter off the excess zinc.
3. Heat the filtrate to evaporate most of the water until it is saturated.
4. Allow the saturated solution to cool for zinc sulfate to crystallise.
5. Filter off crystals.
6. Wash the crystals with cold distilled water. Dry the crystals with filter paper.

E.g.
acid + excess insoluble carbonate
$CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + CO_2(g) + H_2O(l)$

The steps and apparatus are the same as those described for acids with reactive metals except that an insoluble carbonate is used.

The only difference is carbon dioxide is produced instead of hydrogen gas.

This method is used to produce salts of unreactive metals.

E.g.
acid + excess insoluble base
$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$

1. Add excess copper(II) oxide to dilute sulfuric acid, until no more copper(II) oxide dissolves.
2. Filter off the excess copper(II) oxide.
3. Heat the filtrate to evaporate most of the water until it is saturated.
4. Allow the saturated solution to cool for copper(II) sulfate to crystallise.
5. Filter off crystals.
6. Wash the crystals with cold distilled water. Dry the crystals with filter paper.

How are SPA soluble salts prepared?

- **Titration**
  soluble + soluble (acid) → soluble salt

- Sodium, potassium and ammonium do not form insoluble compounds, hence exact volume have to be used to avoid contamination of the salt.
  - acid + alkali
  - acid + carbonate solution

Eg: $HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$

1. Pipette 25.0 cm^3 of aqueous sodium hydroxide into a conical flask. Add 2 – 3 drops of methyl orange. The indicator turns yellow.
2. Add nitric acid slowly from a burette into the conical flask with swirling, until the indicator turns orange permanently. This is the end-point.
3. Note the volume of acid required to neutralise the alkali. Repeat the titration to determine the average volume of acid required.
4. Repeat the titration using the determined volume of acid without an indicator.
5. Heat the sodium nitrate solution until it is saturated. Allow the saturated solution to cool and crystallise.
6. Filter off the crystals. Wash the crystals with cold distilled water. Dry the crystals with filter paper.

Reason for Steps 1 & 2:
The reaction of an aqueous alkali/carbonate solution and acid cannot be determined visually, thus an indicator is necessary to show a colour change which represents complete reaction and no reactant is in excess.

Reason for Step 3:
The amount of each reactant is confirmed and there is no need to add in the indicator anymore, as it will contaminate the salt formed. We want to obtain a pure salt.

Qualitative Analysis (QA)

What is the purpose of adding aqueous sodium hydroxide & aqueous ammonia?

- Dissociate to produce OH^- ions to form an insoluble hydroxide as a precipitate.

How do you remember the results of cation tests?

- First, by colours of ppt. (**coloured ppt. are the most obvious**)
- Next, by the **solubility of the (white) ppt. in excess** aqueous sodium hydroxide & aqueous ammonia; $Zn(OH)_2$ white ppt. dissolves in excess in both aqueous sodium hydroxide & aqueous ammonia.
- Ca^{2+} has very different results compared with the other white precipitates!

Cation	NaOH (aq) Aqueous sodium hydroxide: Add a few drops	NaOH (aq) Aqueous sodium hydroxide: In excess	$NH_3(g) + H_2O(l) \leftrightharpoons NH_4^+(aq) + OH^-(aq)$ Aqueous ammonia: Add a few drops	$NH_3(g) + H_2O(l) \leftrightharpoons NH_4^+(aq) + OH^-(aq)$ Aqueous ammonia: In excess
Ca^{2+} calcium	A white ppt. is formed.	× ppt. is insoluble in excess.	No ppt.	-
Zn^{2+} Zinc	A white ppt. is formed.	✓ ppt. dissolves in excess, forming a colourless solution.	A white ppt. is formed.	✓ ppt. dissolves in excess, forming a colourless solution.
Pb^{2+} Lead(II)	A white ppt. is formed.	✓ ppt. dissolves in excess, forming a colourless solution.	A white ppt. is formed.	× ppt. is insoluble in excess.
Cu^{2+} Copper(II)	A blue ppt. is formed.	× ppt. is insoluble in excess.	A blue ppt. is formed.	✓ ppt. dissolves in excess, giving a dark blue solution.
Fe^{2+} Iron(II)	A dirty-green ppt. is formed.	× ppt. is insoluble in excess.	A dirty-green ppt. is formed.	× ppt. is insoluble in excess.
Fe^{3+} Iron(III)	A reddish-brown ppt. is formed.	× ppt. is insoluble in excess.	A reddish-brown ppt. is formed.	× ppt. is insoluble in excess.
NH_4^+ ammonium	No ppt. Ammonia gas produced on warming.	-	-	-
Na^+, K^+	No ppt. formed.	-	No ppt. formed.	-

What are the identities of the ppt.?

- **Hydroxides of the cations**
- $Ca(OH)_2$, $Zn(OH)_2$, $Pb(OH)_2$, $Cu(OH)_2$, $Fe(OH)_2$, $Fe(OH)_3$
- E.g.

  Chemical equation: $CuSO_4\ (aq) + 2NaOH\ (aq) \rightarrow Cu(OH)_2\ \mathbf{(s)} + Na_2SO_4\ (aq)$

  Ionic equation: $Cu^{2+}\ (aq) + 2OH^-\ (aq) \rightarrow Cu(OH)_2\ \mathbf{(s)}$ **[blue ppt.]**

Why does the ammonium ion produce ammonia gas upon warming with aqueous sodium hydroxide?

- Recall: ammonium salt + alkali → salt + water + ammonia gas

How are anion tests different from cation tests?

- Every single anion has its own separate test reagent(s) while all cations are added with either aqueous sodium hydroxide or aqueous ammonia.
- **Anions are identified either by the production of gas or by the formation of ppt.** (ionic equation are provided in the table below).

Anion	Test	Results
Carbonate, CO_3^{2-}	Add dilute nitric acid. 	Effervescence/Bubbles of gas is observed. **CO_2 gas** forms a **white ppt.** with limewater. $CO_2\,(g) + Ca(OH)_2\,(aq) \rightarrow \mathbf{CaCO_3\,(s)} + H_2O\,(l)$
Nitrate, NO_3^-	Add aqueous sodium hydroxide, and then add **aluminium foil.** Warm.	**Ammonia gas** produced turns moist red litmus paper blue.
Chloride, Cl^-	Acidify with dilute nitric acid. Add aqueous **silver nitrate**.	A **white ppt.** is formed. $Ag^+\,(aq) + Cl^-\,(aq) \rightarrow \mathbf{AgCl\,(s)}$
Sulfate, SO_4^{2-}	Acidify with dilute nitric acid. Add aqueous **barium nitrate.**	A **white ppt.** is formed. $Ba^{2+}\,(aq) + SO_4^{2-}\,(aq) \rightarrow \mathbf{BaSO_4\,(s)}$

Gas	Test	Results
Hydrogen (H_2)	Place a **lighted splint** at the mouth of the test tube.	The lighted splint extinguished with a 'pop' sound.
Oxygen (O_2)	Insert a **glowing splint** into the test tube.	The gas relights the glowing splint.
Carbon dioxide (CO_2)	Bubble the gas into **limewater**, $Ca(OH)_2$.	A white precipitate ($CaCO_3$) is formed in limewater.
Ammonia (NH_3)	Place a **moist red litmus paper** at the mouth of the test tube.	The moist red litmus paper turns blue.
Sulfur dioxide (SO_2) reducing agent, test with oxidising agent	Placed a filter paper soaked with **acidified potassium manganate(VII)** at the mouth of the test tube. **OR** Bubble the gas into acidified potassium manganate(VII).	The purple acidified potassium manganate(VII) turns colourless.
Chlorine (Cl_2) greenish-yellow gas; the rest are colourless	Place a **moist blue litmus paper** at the mouth of the test tube.	Moist blue litmus paper turns red, then is bleached.

Redox

A redox reaction: ***a chemical reaction which involves the oxidation of a reactant and reduction of another reactant <u>at the same time</u>***.

Definitions of oxidation and reduction for reactants:

	Oxidation	Reduction
O atom(s)	gains	loses
H atom(s)	loses	gains
electron(s)	loses	gains
oxidation state	increases	decreases

Oxidising agent: a reactant that causes the other reactant to be oxidised

- giving oxygen; removing hydrogen; or accepting electrons;
- **itself is reduced in the process**

Reducing agent: a reactant that causes the other reactant to be reduced

- removing oxygen; giving hydrogen; or donating electrons
- **itself is oxidised in the process**

Oxidising agents	Reducing agents
oxygen (O_2)	*reactive metals*
chlorine (Cl_2)	*carbon (C)*
bromine (Br_2)	*carbon monoxide (CO)*
*[QA]: **potassium manganate(VII) ($KMnO_4$)***	*[QA]: sulfur dioxide (SO_2)*
	potassium iodide (KI)
	hydrogen (H_2)

1. Redox in terms of O atom(s)

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
Mg: oxidised, **RA**; O_2: reduced, **OA**

$Zn(s) + CuO(s) \rightarrow ZnO(s) + Cu(s)$
Zn: oxidised, **RA**; CuO: reduced, **OA**

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$
Fe_2O_3: reduced, **OA**; CO: oxidised, **RA**

2. Redox in terms of H atom(s)

$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
N_2: reduced, **OA**; H_2: oxidised, **RA**

$H_2S(g) + Cl_2(g) \rightarrow 2HCl(g) + S(s)$
H_2S: oxidised, **RA**; Cl_2: reduced, **OA**

3. Redox in terms of electrons

$Mg \rightarrow Mg^{2+} + 2e^-$ (RA – oxidised)
$Fe^{2+} \rightarrow Fe^{3+} + e^-$ (RA – oxidised)
$Cu^{2+} + 2e^- \rightarrow Cu$ (OA – reduced)
$Cl_2 + 2e^- \rightarrow 2Cl^-$ (OA – reduced)

e.g. metal displacement
$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$
Cu(s): oxidised, lose $2e^-$, **RA**; Ag^+(aq): reduced, each gain $1e^-$, **OA**

e.g. halogen displacement
$Cl_2(aq) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(aq)$

$Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$
Cl_2(aq): reduced, each gain $1e^-$, **OA**; I^-(aq): oxidised, each gain $1e^-$, **RA**

4. Redox in terms of oxidation state

Oxidation state: *the charge an atom of an element would have if it existed as an ion in a compound.*

- This is true even if the atom is covalently bonded.
- When a substance loses or gains electrons, its oxidation state changes

Rules to work out oxidation states:

i. The **oxidation state of a free element is 0.** (Cu, Mg, H_2, O_2 etc)

ii. The **oxidation state of a simple ion is the same as the charge** on the ion. (Cl^-; O.S. is −1 for chlorine)

iii. Some **elements have a fixed oxidation state** in their compounds (Group I is +1, Group II is +2, O is −2, H is +1).

Some possible exceptions:
NaH, sodium hydride; H is -1
H_2O_2, hydrogen peroxide; O is -1
HOCl, hypochlorous acid; Cl is +1

iv. The **oxidation states of all the atoms or ions in a compound add up to 0.**
e.g. CaO; (+2) + (−2) = 0

v. The sum of oxidation numbers of all the atoms in a polyatomic ion is the same as the charge on the ion.
e.g. OH^-; (−2) + (+1) = −1

$$Mg\,(s) + 2HCl\,(aq) \rightarrow MgCl_2\,(aq) + H_2\,(g)$$

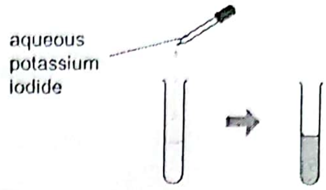
Mg (s)	2HCl (aq)		MgCl₂ (aq)		H₂ (g)
0	+1	−1	+2	2(−1)	2(0)
oxidised	reduced				
RA	**OA**				

Some elements e.g. transition elements can have a variable oxidation state in different substances:

- MnO_4^-, purple: +7
- MnO_4^{3-}, blue: + 5
- MnO_4^{2-}, green: +6
- Mn^{3+}, orange: +3

Not all reactions are redox reactions e.g. neutralisation and precipitation as there is no change in the oxidation states of the elements from reactants to products.

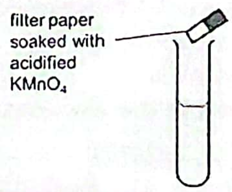
Test for oxidising agents:
Add 2 cm³ of the **reducing agent, aqueous potassium iodide, KI**, to the unknown solution in a test tube.



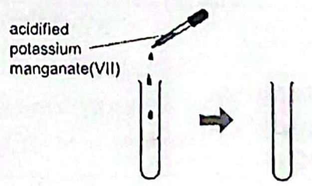
Observation:
Colourless potassium iodide solution turns brown.
Explanation:
Iodide ions are oxidised to iodine molecules by the oxidising agent; the iodine solution formed is brown.

$$2I^-\,(aq) \rightarrow I_2\,(aq) + 2e^-$$
colourless brown

Test for reducing gases:
Place a piece of filter paper soaked with acidified potassium manganate(VII), $KMnO_4$, at the mouth of the test tube.



Test for reducing solutions:
Add 2 cm³ of the oxidising agent, acidified potassium manganate(VII), $KMnO_4$, to the unknown solution in a test tube.



Observation:
Purple acidified potassium manganate(VII) solution turns colourless.
Explanation:

- Manganate(VII) ion MnO_4^- is reduced to manganese(II) ion Mn^{2+}.
- Oxidation state of Mn decreases from +7 to +2.

$$MnO_4^-\,(aq) + 8H^+\,(aq) + 5e^- \rightarrow Mn^{2+}\,(aq) + 4H_2O\,(l)$$
purple colourless

Question: $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$

- Explain why this is a redox reaction?
- How is chlorine acting as an oxidizing agent?

Metals

General Physical Properties of Metals

- good conductors of heat and electricity
- have high melting points and boiling points and densities (except Group I metals)
  - solids at room temperature and pressure (except mercury)
- are malleable and ductile

Diagrams of Metallic and Alloy Structure

Pure Metal — Alloy

Why are pure metals not widely used?

1. Pure metals are soft as the atoms are packed regularly in layers. The layers can slide over one another easily when a force is applied.
2. Pure metals may react with air and water and thus wear away or corrode easily.

Alloys

- **A mixture of a metal with one or a few other elements**
- E.g. brass (Cu & Zn), stainless steel (Fe, C, Cr & Ni)

Why are metals used in the form of alloys?

- **<u>Make metals harder and stronger</u>**
- Improve the appearance of metals
- Lower the melting points of metals
- Make metals more resistant to corrosion

Why are alloys stronger and harder than their constituent metals?

- **The atoms of the different metals or elements have different sizes.**
- **This disrupts the orderly arrangement of atoms in the pure metal, thus making it difficult for the layers of different-sized atoms to slide over each other when a force is applied.**

Metal Reactivity Series

- Arrangement of metals from the most reactive to the least reactive.
- Determined by:
  1. reaction of metals with cold water or steam
  2. reaction of metals with dilute hydrochloric acid
- *More reactive metals have a higher tendency to lose electrons (more reducing) to form positive ions compared to metals that are less reactive.*

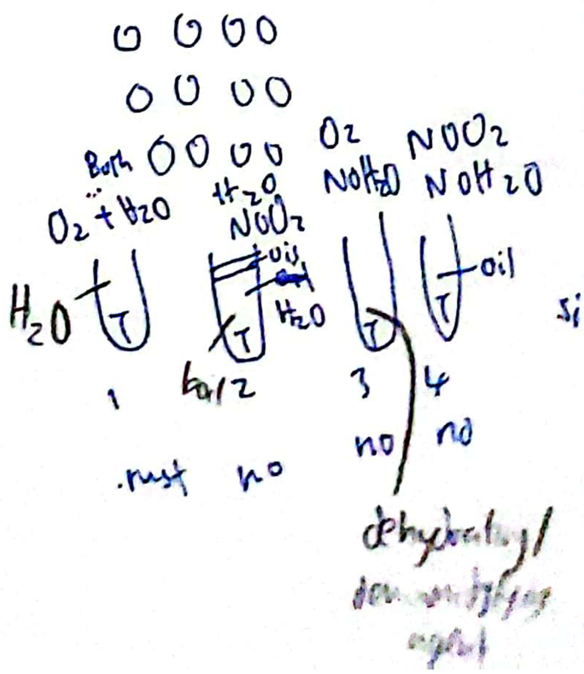
most reactive ↑

Potassium	K
Sodium	Na
Calcium	Ca
Magnesium	Mg
Carbon	**C**
Zinc	Zn
Iron	Fe
Lead	Pb
Hydrogen	**H**
Copper	Cu
Silver	Ag
Gold	Au

least reactive

The Reactivity Series can be used to:

- predict the behaviour of a metal from its position in the reactivity series;
- predict the position of an unfamiliar metal in the reactivity series from a given set of experimental results.



Metal	**Reaction with Cold Water** metal + cold water → metal hydroxide **(aq)** + hydrogen	**Reaction with Steam** metal + steam → metal oxide **(s)** + hydrogen	**Reaction with Hydrochloric Acid** metal + HC*l* (aq) → metal chloride + hydrogen
Potassium	Reacts **very violently**. Enough heat is produced to ignite the hydrogen gas produced. The **hydrogen gas catches fire and burn with a lilac flame.** $2K (s) + 2H_2O (l) \rightarrow 2KOH (aq) + H_2 (g)$	React **explosively.** The reactions should not be carried out in the lsab.	Reacts **explosively.** The reaction should not be carried out in the lab. $2K (s) + 2HCl (aq) \rightarrow 2KCl (aq) + H_2 (g)$
Sodium	Reacts **violently**. The **hydrogen gas may catch fire** and burn with a yellow flame. $2Na (s) + 2H_2O (l) \rightarrow 2NaOH (aq) + H_2 (g)$		Reacts **explosively.** The reaction should not be carried out in the lab. $2Na (s) + 2HCl (aq) \rightarrow 2NaCl (aq) + H_2 (g)$
Calcium	Reacts **moderately fast.** $Ca (s) + 2H_2O (l) \rightarrow Ca(OH)_2 (aq) + H_2 (g)$		Reacts **violently.** $Ca (s) + 2HCl (aq) \rightarrow CaCl_2 (aq) + H_2 (g)$
Magnesium	Reacts **slowly**. A test tube of hydrogen gas is produced only after a few days. $Mg (s) + 2H_2O (l) \rightarrow Mg(OH)_2 (aq) + H_2 (g)$	Hot magnesium reacts **violently** with steam. A bright white glow is produced during the reaction. $Mg (s) + H_2O (g) \rightarrow MgO (s) + H_2 (g)$	Reacts **rapidly.** $Mg (s) + 2HCl (aq) \rightarrow MgCl_2 (aq) + H_2 (g)$
Zinc	No reaction occurs.	Hot zinc reacts **readily** with steam. Zinc oxide is yellow when hot and white when cold. $Zn (s) + H_2O (g) \rightarrow ZnO (s) + H_2 (g)$	Reacts **moderately fast.** $Zn (s) + 2HCl (aq) \rightarrow ZnCl_2 (aq) + H_2 (g)$
Iron		Red-hot iron reacts **slowly** with steam to form iron(II, III) oxide and hydrogen gas. The iron must be heated constantly in order for the reaction to proceed. $3Fe (s) + 4H_2O (g) \rightarrow Fe_3O_4 (s) + 4H_2 (g)$	Reacts **slowly.** $Fe (s) + 2HCl (aq) \rightarrow FeCl_2 (aq) + H_2 (g)$
Lead		No reaction occurs.	The reaction produces an insoluble lead(II) chloride protective layer which prevents the acid from reacting further with lead metal. $Pb (s) + 2HCl (aq) \rightarrow$ **$PbCl_2 (s)$** $+ H_2 (g)$
Copper			No reaction occurs.
Silver			

Displacement Reactions of Metals

- A more reactive metal can displace the ions of a less reactive metal from its salt solution.
- E.g. 1: Place a 2-cm length magnesium ribbon into copper(II) sulfate solution in a test-tube.

$Mg\ (s) + CuSO_4\ (aq) \rightarrow Cu\ (s) + MgSO_4\ (aq)$

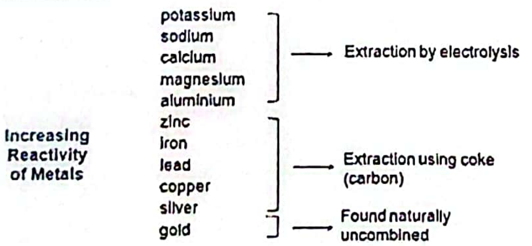
  - ➢ Blue solution starts to fade.
  - ➢ Copper metal is deposited as a pink/ reddish-brown solid on magnesium.

- E.g. 2: Place a copper foil into iron(II) sulfate solution in a test-tube.
  - ➢ No displacement reaction occurs thus there is no visible reaction.

Extraction of Metals

- Metal ore: an ore contains a compound of the metal (oxides, sulfides, chlorides or carbonates) with sand
- The process of obtaining metals from ores is called the extraction of metals.
- There are two main methods to extract metals from their ores. They are:
  - o **Electrolysis of the molten metal compound**, &
  - o **Reduction by carbon** in a blast furnace

The method used to extract a metal depends on the position of the metal in the reactivity series. The methods used for extracting some common metals are summarised below.



- The <u>**position**</u> of the metal in the Reactivity Series determines the method used for its extraction. The more reactive the metal is, the harder it is to extract the metal from its ore.

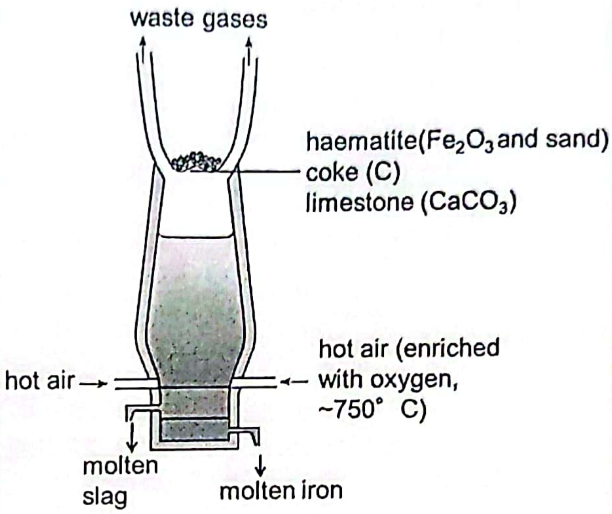
Why is this so?

- **The more reactive the metal, the stronger the electrostatic forces of attraction between the oppositely-charged ions of its compounds.**
- **Hence, more energy is needed to extract the more reactive metals from their compounds.**

Extraction of Iron in a Blast Furnace

Blast Furnace:

Tall, chimney-like; made of bricks; lined with a refractory material (aluminium oxide, magnesium oxide – high melting points and trap heat)



1. Production of carbon dioxide

Coke burns in hot air to produce carbon dioxide and a lot of heat.

$C\ (s) + O_2\ (g) \rightarrow CO_2\ (g)$

2. Production of carbon monoxide

Carbon dioxide rises and reacts with more hot coke to produce carbon monoxide.

$CO_2\ (g) + C\ (s) \rightarrow 2CO\ (g)$

3. Reduction of haematite to iron

Carbon monoxide reduces the iron(III) oxide to molten iron and carbon dioxide.

$Fe_2O_3\ (s) + 3CO\ (g) \rightarrow 2Fe\ (l) + 3CO_2\ (g)$

Molten iron runs to the bottom of the furnace. Hot **waste gases (CO, CO_2 and N_2)** escape through the top of the furnace.

4. Removal of impurities

- Limestone is decomposed by heat to produce carbon dioxide and calcium oxide.
  $CaCO_3\ (s) \rightarrow CaO\ (s) + CO_2\ (g)$

- The **basic oxide**, calcium oxide, reacts with the **acidic impurities**, sand – SiO_2, in the iron ore to form molten slag.
  $CaO\ (s) + SiO_2\ (s) \rightarrow CaSiO_3\ (l)$

Lighter slag runs down and floats **on top** of the molten iron. Molten slag and molten iron can be tapped separately from the furnace.

Rusting of Iron

- The corrosion of iron.
- Essential conditions for rusting: presence of both oxygen and water
- Rust – the hydrated iron(III) oxide, $Fe_2O_3.xH_2O$

Rust Prevention Methods

1 **Using a protective layer**
To keep iron/ steel from coming into contact with oxygen and water
  - E.g. painting, greasing, covering with plastic, electroplating – tin or zinc
  - This layer cannot be scratched or damaged.

The Need to Recycle Metals

- Metals are **finite resources** and will run out one day. Recycling metals will help to **conserve metals – to ensure that the supply of metals last longer.**

Recycling Steps

- Metal objects which are no longer needed are collected,
- crushed and
- melted into blocks of clean metals.

Advantages of Recycling Metals

1 Better conservation of natural resources such as fossil fuels as extraction of metals from their ores requires a continuous supply of energy; by recycling metals, there is no need for metal extraction.

2 Reduce environmental problems related to extracting metals from their ores and better land use.
  - Mined land cannot support life.
  - Waste from extraction, if not disposed properly, will leach into soil and pollute the land and water.
  - Recycling frees up land for agriculture and housing and greatly reduce pollution.

3 Saves cost of extracting metals from their ores as fewer landfills need to be built to dispose used metal objects and waste material from metal extraction.

Economic, Social and Environmental Issues of Recycling Metals

- Recycling can be extremely costly, and sometimes more expensive than metal extraction – for transport, separating and cleaning different types of metals. Recycling will not be worthwhile if the costs are too high and the metal recycled is not an expensive or valuable one.
- It takes time and effort for communities to adopt recycling as a lifestyle.
- Recycling may cause pollution. E.g. recycling of lead by melting car batteries releases harmful gases, causing air pollution.

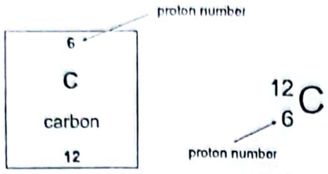
The Periodic Table

How are elements arranged in the Periodic Table?

- elements are arranged in order of *increasing proton (atomic) number*

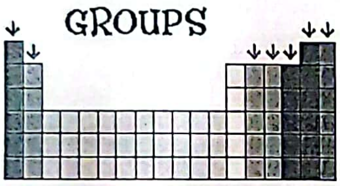
Position of an Element

- proton number → electronic structure of an element

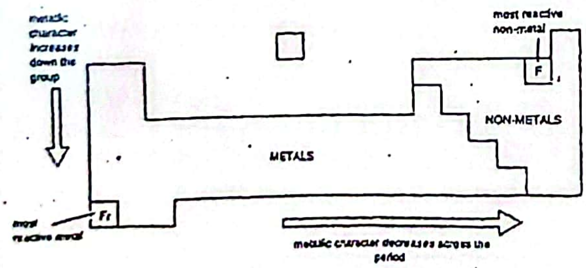


- e.g. carbon has a proton number **6**
  - electronic structure: **2.4**
  - belongs **group IV** due to C having four valence electrons
  - belongs to **Period 2** due to C having two electron shells.

Groups



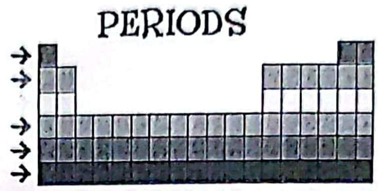
- eight groups of elements; groups run from top to bottom (vertical columns)
- ***number of valence electrons is the same* as the group number of the element**
  - elements in the same group will *have the same number of valence electrons* and *similar chemical properties*



- The **metals** are found on the left-hand side of the Table and are elements with 1 – 3 valence electron(s).
- The **non-metals** are found on the right-hand side of the Periodic Table and are elements with 4 – 7 valence electron(s).
- The **noble gases/ Inert gases** are non-metals that are found on the extreme right-hand side, and are elements with a fully-filled valence electron shell of 2 (He) or 8 valence electrons.

- Going down a group, there is an ***increase in metallic properties* and a *decrease in non-metallic properties.***
  - As the size of the atom increases, the valence electrons will be further away from the attractive forces of the positively-charged nucleus (due to presence of protons);
  - An element further down a group will lose its electron(s) more easily.

Periods



- seven periods; periods run from left to right (horizontal rows)
- ***number of electron shell(s) is the same* as the period number of the element**

Periodic <u>Trends</u>:

- elements change from metals to metalloid to non-metals across a period
  - **there is a *decrease in metallic properties* and an *increase in non-metallic properties*: a stable fully-filled valence shell is reached not from losing electrons but from gaining electrons**
  - **the nature of their *oxides changes from <u>basic to amphoteric to acidic</u>* correspondingly**

<u>Group I: Alkali Metals</u>

Physical Properties

- **soft, can be cut easily by a knife**
- have **low melting and boiling points**
- have **low densities** and the first 3 alkali metals float on water

Physical <u>Trends</u> Down Group I

- melting points decrease (Rubidium m.p. 39 °C, b.p. 695 °C if asked to predict)
- densities generally increases

Notable Properties

- highly reactive metals that are ***stored in oil***
- having 1 valence electron, alkali metals tend to lose this electron to form ions of 1+ charge and their oxidation is +1.
  - e.g. $K \rightarrow K^+ + e^-$ (redox)
  - *powerful reducing agents* in all their reactions, *forming ionic compounds*

Chemical Properties:

- **reacts with cold water to form an alkali and hydrogen**

  Place 2 g of metal into a beaker containing 100 cm^3 cold water.

  Li: reacts quickly. Lithium floats on water. No flame is seen.
  $2Li\ (s) + 2H_2O\ (l) \rightarrow 2LiOH\ (aq) + H_2\ (g)$

  Na: reacts violently. Sodium melts into a silvery ball and darts about on the surface of water. The hydrogen gas may catch fire and burn with a yellow flame.
  $2Na\ (s) + 2H_2O\ (l) \rightarrow 2NaOH\ (aq) + H_2\ (g)$

  K: reacts very violently. Potassium burst into flames upon contact with water. The hydrogen gas may catch fire and burn with a lilac flame.
  $2K\ (s) + 2H_2O\ (l) \rightarrow 2KOH\ (aq) + H_2\ (g)$

Chemical Trend Down Group I

- reactivity increases (Francium is the most reactive)
  - as the size of the atom increases, it becomes easier to lose its valence electron as this electron is further away from the positively-charged nucleus.

Group VII: Halogens

Physical Properties

- exist as diatomic molecules (F_2, Cl_2, Br_2, I_2 and At_2) with **low melting and boiling points**
  - a low amount of energy is needed to overcome the weak intermolecular forces of attraction between halogen molecules
  - coloured

halogen	state at r.t.p	colour
Cl_2	gas	greenish-yellow
Br_2	liquid	reddish-brown
I_2	solid	purplish-black
At_2	solid	black

Physical Trends Down Group VII

- melting points and boiling points increase (physical state changes from gases to liquid to solids)
- colours of halogens become darker

Notable Properties

- highly reactive non-metals
  - having 7 valence electrons, each halogen atom tends to gain one electron to form ions of 1− charge. e.g. $Cl_2 + 2e^- \rightarrow 2Cl^-$
  - *powerful oxidising agents* in all their reactions, *forming halides* (F^-, Cl^-, Br^- and I^- ions)

Chemical Properties

- **Displacement reaction:** a more reactive halogen will displace a less reactive halogen from its halide solution.
  (IMPORTANT!)

How to determine the reactivity of halogens?

- **E.g.** Place 2 cm^3 of aqueous chlorine into and equal volume of potassium bromide solution in a test-tube.
- Repeat with other halogens and halide solutions.

halide solution / halogen	KCl (aq)	KBr (aq)	KI (aq)
Cl_2 (aq)		✓	✓
Br_2 (aq)	×		✓
I_2 (aq)	×	×	

Observations:

1. $Cl_2\ (aq) + 2KBr\ (aq) \rightarrow 2KCl\ (aq) + Br_2\ (aq)$
   - colourless potassium bromide solution turned **reddish-brown** due to the formation of **aqueous bromine** (when light yellow aqueous chlorine was added)

2. $Cl_2\ (aq) + 2KI\ (aq) \rightarrow 2KCl\ (aq) + I_2\ (aq)$
   - colourless potassium iodide solution turned **brown** due to the formation of **aqueous iodine** (when light yellow aqueous chlorine was added)

3. $Br_2\ (aq) + 2KI\ (aq) \rightarrow 2KBr\ (aq) + I_2\ (aq)$
   - colourless potassium iodide solution turned **brown** due to the formation of **aqueous iodine** (when reddish-brown bromine water was added)

Chemical Trend Down Group VII

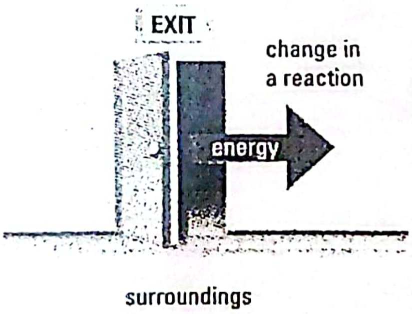
- reactivity decreases
  (Fluorine most reactive)
  - as the size of the atom increases, it becomes more difficult for the nucleus to attract one more electron electrostatically as the valence shell is further away. i.e. larger halogens do not gain the electron as easily as smaller ones.

Group 0: Noble Gases

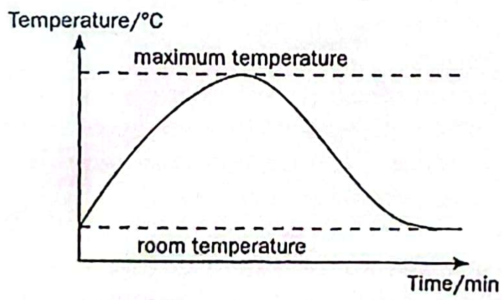
- monatomic colourless gases
- described as stable, unreactive or inert
  - **Noble gases have a *fully-filled valence electron shell*. They do not need to lose, gain or share electrons with other elements to form compounds.**
- Uses of some noble gases:
  - He: fill weather balloons/airships
  - Ar: fill tungsten bulbs to **provide an *inert* environment to prevent oxidation** of the filament or provide inert environment in the manufacture of steel and welding stainless steel
  - Ne: fill tungsten bulbs or making lights and advertising signs

Exothermic change

- a process or chemical reaction which **gives off heat to the surroundings** and this may be detected by an **increase in temperature of the surroundings**



Variation of temperature with time for an exothermic reaction:

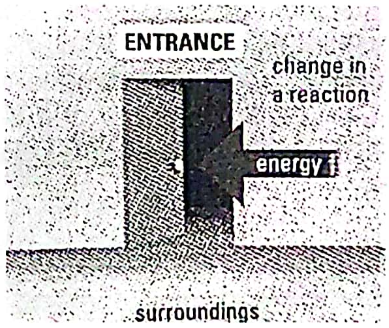


Examples of exothermic reactions:

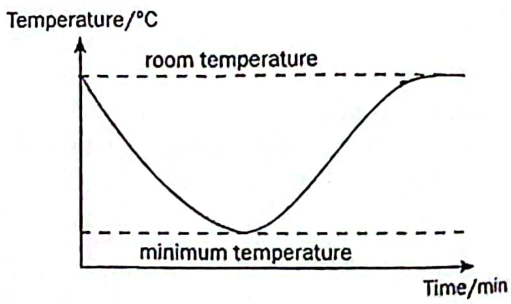
- State change: **condensation & freezing**
- **Combustion** reaction
- **Neutralisation** reaction (eg: HCl + NaOH)
- Metal displacement reaction

Endothermic change

- a process or chemical reaction which **takes in heat from the surroundings** and this may be detected by a **decrease in temperature of the surroundings**



Variation of temperature with time for an endothermic reaction:



Examples of endothermic reactions:

- State change: **melting & boiling**
- Thermal decomposition of compounds
- Reaction between acids and hydrogencarbonates
- **Dissolving of ammonium nitrate in water**

Speed of Reaction

EFFECTIVE COLLISIONS = REACTION

For a **reaction to occur**, the reacting particles must

- **collide** with each other in the correct orientation, and
- **collide with enough energy** (the *activation energy*) to break the existing bonds.

The higher the frequency of effective collisions, the faster the speed of reaction.

SPEED OF REACTION

- defined as how fast a reaction progresses over an interval of time.
- can be measured as:
- **the amount of a reactant used up <u>per unit time</u>. (e.g. mass loss due to gas produced per unit time)**
- **the amount of a product formed <u>per unit time</u>. (e.g. volume of gas produced per unit time)**

HOW DO WE MEASURE SPEED OF REACTION IN THE LABORATORY?

➢ measured by some **<u>detectable changes with time</u>**

- **Measuring Speed from <u>Changes in Gas Volume</u>**

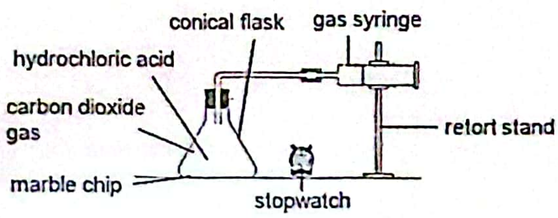
  If **one of the products of the reaction is a gas**, the speed of the reaction can be measured as the **volume of gas produced per unit time.**

  What type of reaction can this method be used for?

- **Carbonate + acid → salt + *<u>carbon dioxide</u>* + water**
- **Metal + acid → salt + *<u>hydrogen</u>***

For example:

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



1 Add 1 g of calcium carbonate to 50 cm^3 of dilute hydrochloric acid (1 mol/dm^3) **in a 250 cm^3 conical flask.** When carbonates react with acids, a salt and carbon dioxide gas is produced.

2 Close the conical flask with the stopper of a **delivery tube** that is connected to a 500 cm^3 **gas syringe** via a rubber tubing.

3 Record the **volume of carbon dioxide produced at fixed time intervals of 30 seconds.**

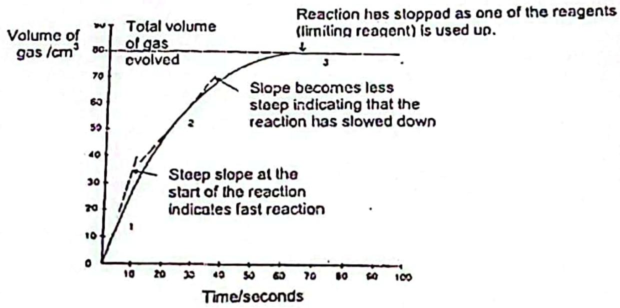
4 Stop recording the volume of carbon dioxide produced after the reaction stops and the reading from the gas syringe no longer changes.

5 **Plot a graph** of volume of carbon dioxide against the time taken for the reaction to be completed.

6 The **speed** of the reaction at any time can be determined by finding the **gradient of the graph.** The **average speed** is found by dividing the final total volume of carbon dioxide produced by the time taken

7 and the **units is cm^3/s.**

The **gradient of graph** indicates the **speed of reaction:**



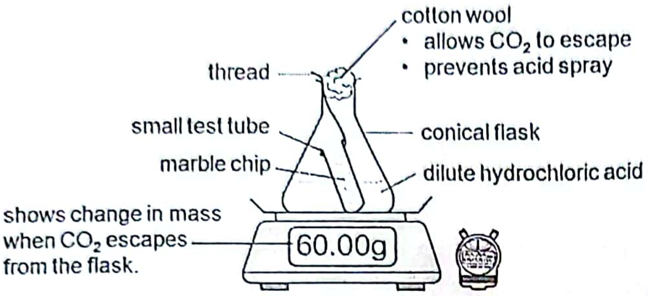
HOW DO WE EXPLAIN THE GENERAL SHAPE OF RATE GRAPHS?

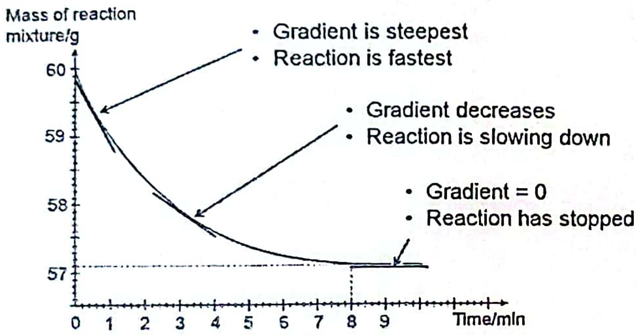
➢ The **speed** of the reaction is the **highest at the beginning.** The gradient of the graph is the largest as the **concentration of reactants is the greatest, highest frequency of effective collisions.** The steeper the gradient, the faster the reaction.

➢ The **speed** of the reaction **decreases with time** as the **concentration of reactants decreases, lower frequency of effective collisions.** The gradient of the graph decreases.

➢ The reaction has **stopped and no more products are produced when one (limiting reagent) or both of the reactants are fully reacted.** The graph is horizontal and gradient is 0.

- **Measuring Speed from <u>the Decrease in Mass of Reactant at Regular Time Intervals</u>**
  We can also measure the decrease in mass of the reaction mixture **if one of the products is a gas**. When the gas <u>escapes to the surrounding</u>, there will be a decrease in the mass of the flask and its contents.





FACTORS AFFECTING THE SPEED OF CHEMICAL REACTION

1. **Pressure** (*gaseous* reactants only)

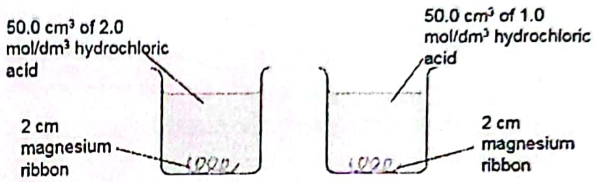
- The **higher the pressure, the more the <u>reacting gaseous particles per unit volume</u>**.
- The reacting gaseous particles are closer together, hence the **<u>frequency of effective collisions increases</u>**, more product formed, thus the speed of reaction increases.
- **The higher the pressure of <u>gaseous</u> reactants, the faster the speed of reaction.**

2. **Concentration** (of *solutions e.g. acids*)

- The **more concentrated** the solution, **the <u>more reacting particles per unit volume</u>**.
- The reacting particles are closer together, hence the **<u>frequency of effective collisions increases</u>**, more product formed, thus the speed of reaction increases.

➢ **The higher the concentration of the reactants, the faster the speed of reaction.**

Experiment to Investigate Factor:

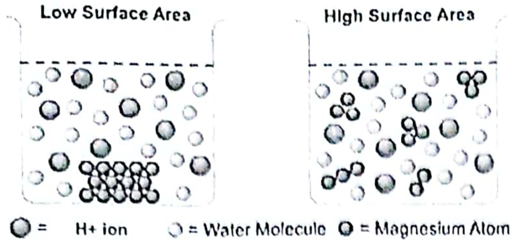


1. Using a measuring cylinder, measure 50 cm^3 of 1 mol/dm^3 HCl in a beaker.
2. Add 2 cm of magnesium ribbon into the HCl in the beaker and start the stopwatch.
3. As soon as the magnesium has completely dissolved, stop the stopwatch and record the time taken.
4. Repeat steps 1 – 3 with 50 cm^3 of 2 mol/dm^3 HCl using fresh magnesium (2 cm).
5. Record the time taken in a suitable format e.g. (table).

conc. of HCl / mol/dm^3	1	2
time taken/s		

The shorter the time taken, the faster the speed of reaction.

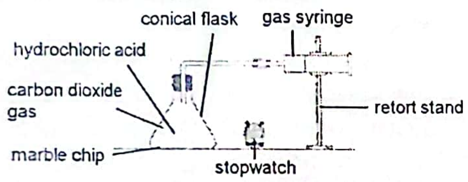
3. Surface area/particle size



(Surface area ∝ 1 / particle size)

- The **smaller** the solid reactant **particle size**, the **larger the exposed surface area** of these particles.
- There is more exposed surface area for the other reactant particles to collide into.
- Therefore, the **frequency of effective collisions increases**, more product formed, thus the speed of reaction increases.
- **The smaller the particle size, the larger the exposed surface area, the faster the speed of reaction.**

Experiment to Investigate Factor:



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

- Measure 2 g of $CaCO_3$ **powder** into a conical flask.
- Add 100 cm³ of 1 mol/dm³ HC*l* into the flask. The CO_2 produced will be collected by the gas syringe.
- Start the stopwatch immediately and stop the timing when **50 cm³ of CO_2 is produced.**
- Repeat steps 1 – 3 using 2 g of $CaCO_3$ **lumps.**

The larger the exposed surface area of calcium carbonate (the smaller the particle size of $CaCO_3$), the shorter the time taken by the reaction, the faster the speed of reaction.

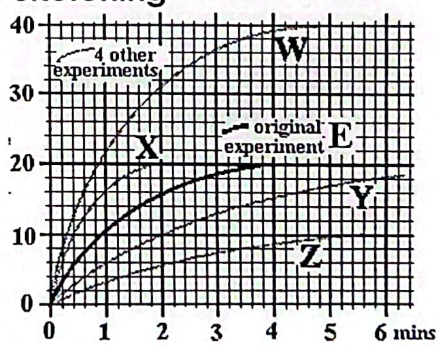
4. Temperature

- As the **temperature increases**, more particles have energy greater than the activation energy and the **reacting particles gain more energy, moving faster**.
- The **frequency of effective collisions increases**, more product formed, thus the speed of reaction increases.
- **The higher the temperature, the faster the speed of reaction.**

Experiment to Investigate Factor:

1. 50 cm³ of HC*l* is measured and placed in a small beaker.
2. Add a strip of magnesium ribbon (2 cm) into the small beaker with HC*l* at 20 °C.
3. Start the stopwatch immediately and stop the timing when the magnesium ribbon has just fully disappeared. Record the time taken.
4. Repeat steps 1 – 3 at **different temperatures** of 30 °C, 40 °C, 50 °C and 60 °C of HC*l*.
5. The higher the temperature, the shorter the time taken by the reaction, the faster the speed of reaction.

Curve Sketching



Graph E: Original experiment; initial concentration, initial mass of solid and its particle size and specific initial and constant temperature

Graph W: double the mass of solid / no. of moles of acid (if limiting)

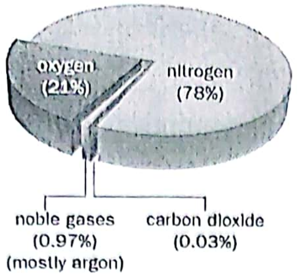
Graph X: higher conc of acid (excess) / higher temperature / smaller particle size of solid e.g. powder

Graph Y: lower conc of acid (excess) / lower temperature / larger particle size of solid e.g. lumps

Graph Z: halving the mass of solid / no. of moles of acid (if limiting)

AIR

Composition of clean air (unless otherwise stated):



Some common atmospheric pollutants:

- carbon monoxide;
- methane;
- nitrogen oxides (NO and NO_2);
- ozone;
- sulfur dioxide;
- unburnt hydrocarbons

Source of **Carbon Monoxide:**

- **Incomplete combustion of carbon-containing fuels** in vehicles or power plants due to insufficient amount of oxygen

Effect of Carbon Monoxide:

- CO **binds irreversibly with haemoglobin** in red blood cells, **reducing the ability of haemoglobin to transport oxygen** to parts of the body, thus **causing death.**

Sources of **Sulfur Dioxide** ($S + O_2 \rightarrow SO_2$):

- Man-made source: Combustion of **sulfur-containing** fossil fuels in power stations and some factories
- Natural source: Volcanic eruptions

Effect of Sulfur Dioxide:

- Dissolves in rain water and is oxidised by oxygen in the air to form **sulfuric acid,** contributing to **acid rain** which **corrodes metals/ limestone in buildings** and **destroys aquatic plants and animals.**

Sources of **Oxides of Nitrogen:**

- Man-made source: Combustion of N_2 at high temperatures in **internal combustion engines** of vehicles due to high temperature
- Natural source: Lightning activity

Effect of Nitrogen Dioxide:

- Dissolves in rain water and is oxidised by oxygen in the air to form **nitric acid,** contributing to **acid rain** which **corrodes metals/ limestone in buildings** and **destroys aquatic plants and animals.**

DID YOU KNOW?
Unpolluted rainwater is slightly acidic (pH slightly below 7) because carbon dioxide from the air dissolves in rainwater to form carbonic acid, which is a weak acid.

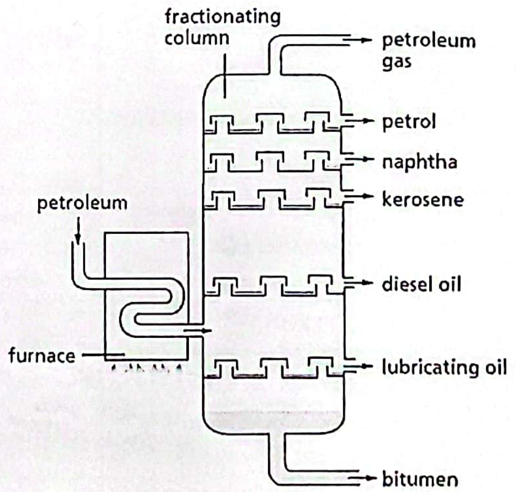
FUELS

Main **sources of fuel** today:

- **petroleum (crude oil)** – mixture of hydrocarbons
- **natural gas (mainly methane)**

Fractional Distillation of Crude Oil:

1. In the furnace at the bottom of the fractionating column, petroleum is heated into a **vapour.**
2. The vapour is pumped into a **huge fractionating column** where the fractions are separated by their **boiling points.**
3. Hot vapour rises up the column and begins to **cool and condense.**
4. The **lighter fractions have lower boiling points** and are **collected at the top** of the fractionating column **as gases** with **petroleum gas as the first fraction.**
5. **Heavier fractions have higher boiling points** and are collected at the **lower sections** of the fractionating column with **bitumen as the residue.**



Some uses of the different fractions:

- **Petroleum gas:** fuel for cooking
- **Petrol (gasoline):** fuel for car engines
- **Naphtha:** feedstock for petrochemical industries
- **Kerosene (paraffin):** fuel for aircraft engines
- **Diesel oil:** fuel for diesel engines
- **Lubricating oil:** lubricants for machines and a source of polishes and waxes
- **Bitumen:** for making road surfaces

Boiling point range of fraction **increases**

e.g. alkene C=C
e.g. Alkane

e.g. alcohol O–H
e.g. carboxylic acid COOH → C(=O)–O–H

What is a **hydrocarbon**?

- A **organic compound** that is **only** made up of **hydrogen** and **carbon** chemically (covalently) combined together.

What is a **homologous series**? → same functional group

- A homologous series is a group of compounds
  - with the same general formula, (where n is an integer)
  - similar chemical properties and
  - show a gradation in physical properties as a result of increase in the size and mass of the molecules. (melting and boiling points, viscosity, density increase while flammability decreases as a result of the increase in molecular size)

1 C: meth–
2 C: eth–
3 C: prop–
4 C: but–

All organic compounds can undergo **combustion** reactions:

Complete combustion
(with sufficient/excess/enough oxygen):

- hydrocarbon (alkane/alkene) + oxygen → carbon dioxide CO_2 **(g)** + water vapour H_2O **(g)**

Incomplete combustion
(with insufficient/limited/not enough oxygen):

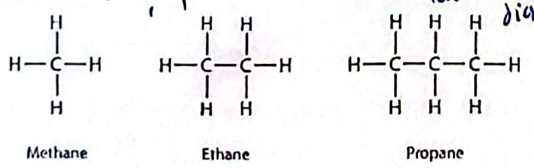
- will produce **carbon monoxide (CO)** + water vapour

→ Everything undergo combustion → use up oxygen

Alkanes (no functional group)

General formula of **alkanes**: □ + O_2 → CO_2 + H_2O (form carbon dioxide, water vapour)

- C_nH_{2n+2} ★ –C–C–



Saturated compound:

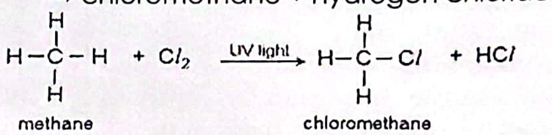
- A compound containing only carbon–carbon single bonds and does not contain carbon-carbon double bonds.

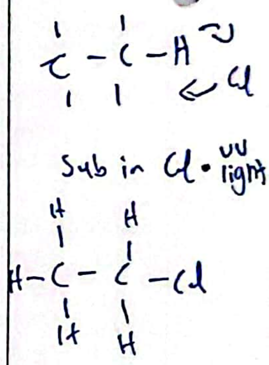
Physical properties of alkanes:

- Insoluble in water, soluble in organic solvents.
- As molecular size increases, **moving up through the members:**
  - m.p & b.p. increase.
    - As the molecular size of alkanes increases, the intermolecular forces of attraction increases.
    - A larger amount of energy is required to overcome these forces of attraction during boiling for a larger alkane.
  - viscosity increases.
  - density increases
  - flammability decreases.

★ **Substitution** reaction of alkanes: → In the football field, you sub out players.

- **Condition: UV light**
- Eg: methane + chlorine → chloromethane + hydrogen chloride





Cracking of Long-chain Alkanes

Conditions for cracking:

- Al_2O_3 and SiO_2 catalyst
- Approximately 600 °C (high temp.)

Why is cracking important?
To **break up longer-chain alkane** molecules into more useful smaller-chain molecules.

- **short-chain alkenes** (making plastics)
- **hydrogen** (making ammonia)
- **short-chain alkanes to produce petrol** as fuel.

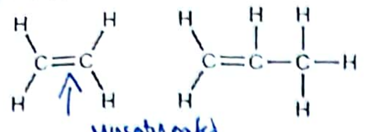
Description:

- Catalytic cracking is done by passing the petroleum fractions containing long-chain alkanes over finely-divided aluminium oxide (Al_2O_3) and silicon dioxide (SiO_2) catalysts at a high temperature of 600 °C.
- The product is a mixture of short-chain alkenes and short-chain alkanes or hydrogen gas.

Alkenes (C=C double bond)

General formula of alkenes:

- C_nH_{2n}



ethene propene

Unsaturated compound:

- Compound containing carbon–carbon double bond(s).
- Atoms can be added across the double bond(s).

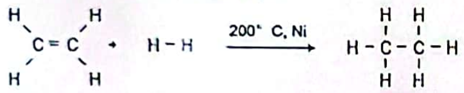
Polyunsaturated compound:

- Compound containing more than one carbon–carbon double bond.

E.g. vegetable oil

Addition reactions of alkenes:

- Addition of hydrogen (**hydrogenation**)
  - **Conditions: Heating (200 °C), nickel catalyst**

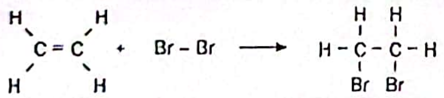


ethene hydrogen ethane

Application: Manufacture of margarine

vegetable oil $\xrightarrow[\text{Ni catalyst}]{\text{hydrogen, 200 °C}}$ margarine

- Addition of **aqueous bromine (bromination)**
  - Condition: room temperature



ethene bromine 1,2-dibromoethane

Application:

****How do we differentiate between alkanes VS alkenes, ethane VS ethene or saturated fats VS unsaturated fats or margarine VS vegetable oil?**

- Test: Add reddish-brown aqueous bromine to both samples in separate test-tubes.
- Observation with saturated compounds (eg: ethane): Aqueous bromine remains **reddish-brown.**
- Observation with unsaturated compounds (eg: ethene): **Reddish-brown aqueous bromine is decolourised.**

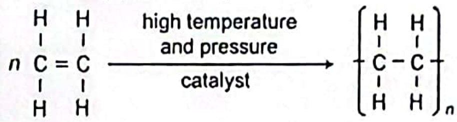
- Addition polymerisation
  - Conditions: high temperature & pressure, catalyst

What is an addition polymer?

It is a macromolecule formed by many similar small molecules (monomers) added across their carbon–carbon double bond.

- Each alkene monomer contains a carbon-carbon double bond and **one bond in the carbon-carbon double bond breaks** in each alkene monomer.
- Each monomer forms **two new carbon-carbon single bonds with two other alkene monomers** and eventually.
- **Many monomers join together** to form the polymer.

e.g. Poly(ethene)



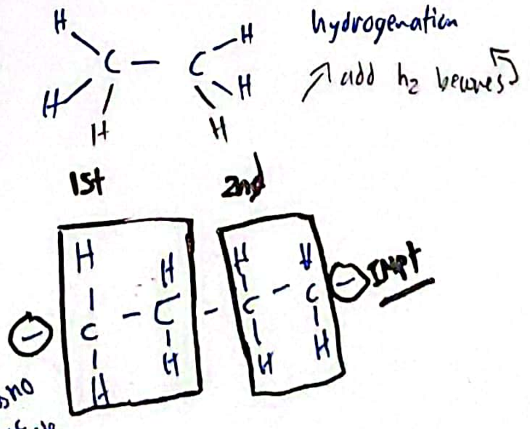
OR

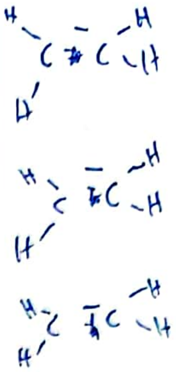
$$n(CH_2{=}CH_2) \xrightarrow[\text{catalyst}]{\text{high temperature and pressure}} \left[CH_2{-}CH_2\right]_n$$

Plastics and Pollution

Improper disposal of plastics can cause air, water and land pollution as plastics are **non-biodegradable** as they **cannot be decomposed by soil bacteria** and will remain for a long time.

- Solution: reuse/ recycle plastics.

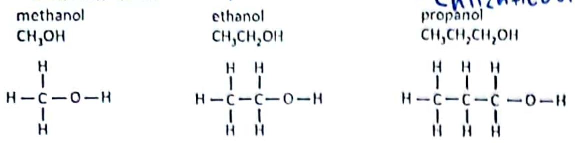




Alcohols (hydroxyl group, –OH)

General formula of **alcohols**: carboxylic acid

- $C_nH_{2n+1}OH$ =7 Oxidise =7 $-C_nH_{2n+1}COOH$



Oxidation reaction of alcohols:

- Condition: heat
- Reagents: **aqueous $KMnO_4$, dilute H_2SO_4**
- Observation: purple $KMnO_4$ turned colourless

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

- Eg: C_2H_5OH (aq) + **2[O]** $\xrightarrow{\text{heat}}$ CH_3COOH (aq) + H_2O (*l*)

Physical properties of alcohols:

- As molecular size increases,
  - solubility in water decreases.
  - b.p. increases.

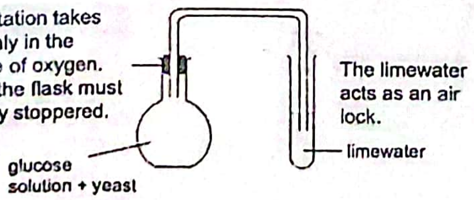
Production of ethanol:

- Hydration of alkene
- **Fermentation of glucose to produce ethanol****

$C_6H_{12}O_6(aq) \xrightarrow{\text{yeast}} 2C_2H_5OH(aq) + 2CO_2(g)$

1. Yeast is added to **glucose solution** in a conical flask at a temperature of **37 °C** in the **absence of oxygen** to **produce ethanol solution and carbon dioxide** gas.
2. **Pure ethanol** can be obtained from the mixture by **fractional distillation**.

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



Carboxylic Acids (carboxyl group, –COOH)

General formula of **carboxylic acids**:

- $C_nH_{2n+1}COOH$

- have acidic properties due to the hydrogen ions produced.
  - ✓ Acid + alkali → salt + water
  - ✓ Acid + (reactive) metal → salt + hydrogen
  - ✓ Acid + carbonate → salt + water + carbon dioxide

Physical properties of carboxylic acids:

- As molecular size increases,
  - solubility in water decreases.

Why does ethanol turn sour over time?

Production of ethanoic acid from ethanol:

- Heating a mixture of ethanol and an oxidising agent:
  - o acidified $KMnO_4$ (colour change: purple to colourless), OR

CH_3CH_2OH (aq) + **2[O]** $\xrightarrow{\text{heat}}$ CH_3COOH (aq) + H_2O (*l*)

- **Oxidation of ethanol by atmospheric oxygen in the presence of certain bacteria**

CH_3CH_2OH (aq) + **O_2 (g)** $\xrightarrow{\text{bacteria}}$ CH_3COOH (aq) + H_2O(*l*)

cannot go beyond 37°C =7 enzymes denatured
cannot oxygen because will oxidise.

☆Ethanol
↳ fermentation – glucose
– yeast

Meth | Mp/bp↑
eth | viscosity↑
prop | density↑
but ↓ | flammability↓

Charges of Ions

Group I	+
Group II	2+
Group III	3+
Group V	3-
Group VI	2-
Group VII	-

Chemical Formula

Ammonium	NH_4^+
Silver	Ag^+
Zinc	Zn^{2+}
Iron (II)	Fe^{2+}
Iron (III)	Fe^{3+}
Hydroxide	OH^-
Nitrate	NO_3^-
Carbonate	CO_3^{2-}
Sulfate	SO_4^{2-}

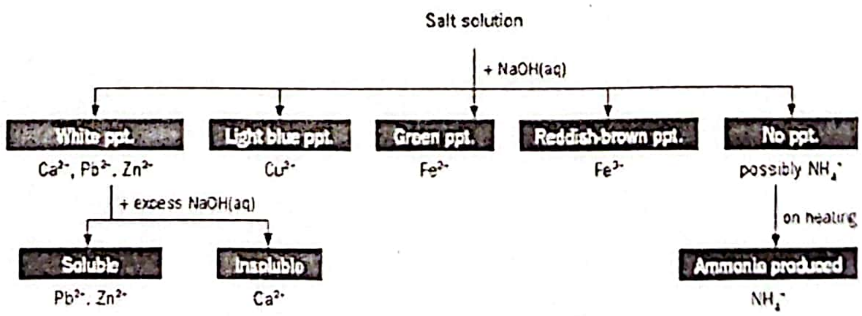
Solubility Table

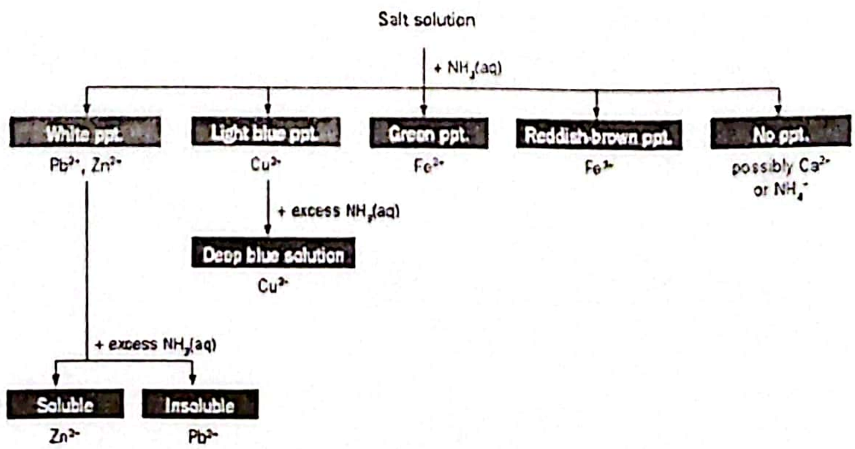
Soluble salts		Insoluble salts
all sodium salts all potassium salts all ammonium salts all nitrates		–
all chlorides	except →	silver chloride, AgCl lead(II) chloride, $PbCl_2$
all sulfates	except →	barium sulfate, $BaSO_4$ lead(II) sulfate, $PbSO_4$ calcium sulfate, $CaSO_4$ (sparingly soluble)
sodium carbonate, Na_2CO_3 potassium carbonate, K_2CO_3 ammonium carbonate, $(NH_4)_2CO_3$	← except	all carbonates

Reactivity Series

(Most reactive) Potassium
Sodium
Calcium
Magnesium
Carbon
Zinc
Iron
Lead
Hydrogen
Copper
Silver
(Least reactive) Gold

Cation Test





The Periodic Table of Elements

Key: atomic number / **atomic symbol** / name / relative atomic mass

1+	2+							Group				3+	4+ 4−	3−	2−	1−	0
I	II											III	IV	V	VI	VII	0
							1 H hydrogen 1										2 He helium 4
3 Li lithium 7	4 Be beryllium 9											5 B boron 11	6 C carbon 12	7 N nitrogen 14	8 O oxygen 16	9 F fluorine 19	10 Ne neon 20
11 Na sodium 23	12 Mg magnesium 24											13 Al aluminium 27	14 Si silicon 28	15 P phosphorus 31	16 S sulfur 32	17 Cl chlorine 35.5	18 Ar argon 40
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	31 Ga gallium 70	32 Ge germanium 73	33 As arsenic 75	34 Se selenium 79	35 Br bromine 80	36 Kr krypton 84
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
actinoids	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 –

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