

CHEMISTRY

FORM THREE

COMPOSED BY:

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TOPIC 1

CHEMICAL EQUATIONS

CHEMICAL EQUATION

- Is the chemical statement which represents a chemical reaction by symbols and formulae.
- It consists of two sides reactant and product side.
- The numbers may be written in front of the formula or symbol, these are called stoichiometric values/coefficients.

Features of a Chemical Equation

- It should have reactant side and product side.
- The atoms that appears in the reactant side must also appear in the product side.
- The number of each kind of atom should be equal in the reactant and product side.

Symbols used in chemical equations

- **Arrow** (\rightarrow or \rightleftharpoons) means “yields” or “produces”. It separates the reactant and product side.
- **Delta** (Δ) means heat is supplied to the reaction. It is put on the reaction arrow($\xrightarrow{\Delta}$).
- **Plus** (+) in reactant means “react with” but in product means “and”. It separates one reagent and the other.
- Physical state symbols,
 - (s) Denotes a solid
 - (l) Denotes a liquid
 - (g) Denotes a gas
 - (aq) Denotes an aqueous solution(something dissolved in water).

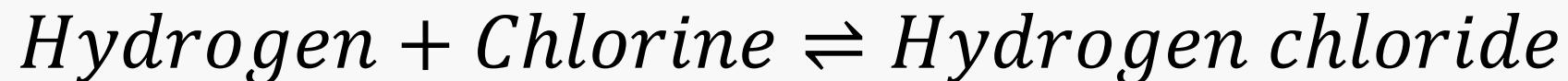
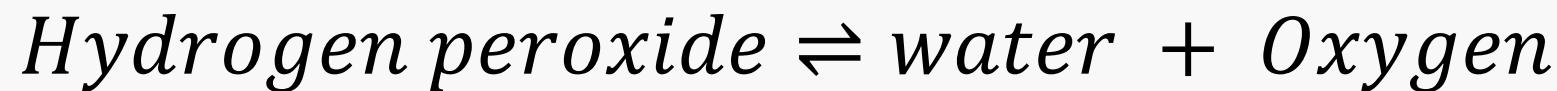
Types of chemical equation

- There are three types of chemical equation include
 - i) Word equation
 - ii) Molecular(formula) equation
 - iii) Ionic equation

i) Word equation

◦ A word equation is the equation which is expressed in words.

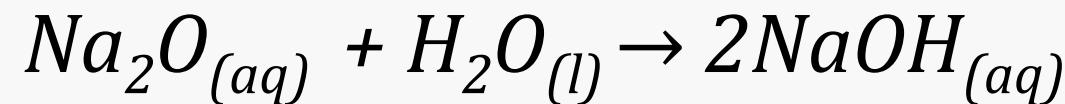
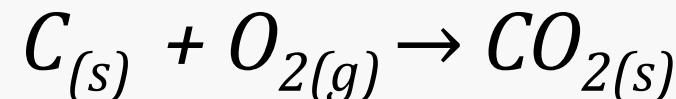
Example:



ii) Molecular (formual) equation

- Molecular equation is the equation which is written using the chemical symbols and formulas.

Example:



Difference between molecular equation and word equation

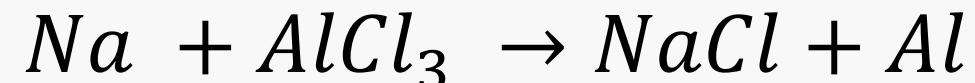
	MOLECULAR EQUATION	WORD EQUATION
1.	It can be balanced	It can't be balanced.
2.	It shows the physical states of the reagents	Does not show the physical states of the reagents
3.	Shows relative proportion of reacting substances	Does not shows relative proportion of reacting substances
4.	It occupies very small space but includes a lot of information.	It occupies large space but includes little information.
5.	The heat changes involved in a reaction are written.	The heat changes involved in a reaction are not written.

Steps for balancing chemical equation

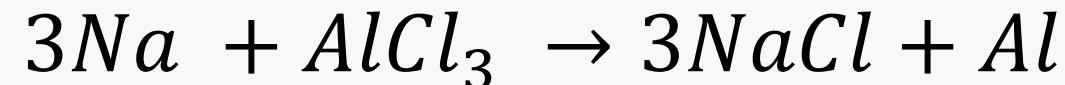
- i. Write equation in word



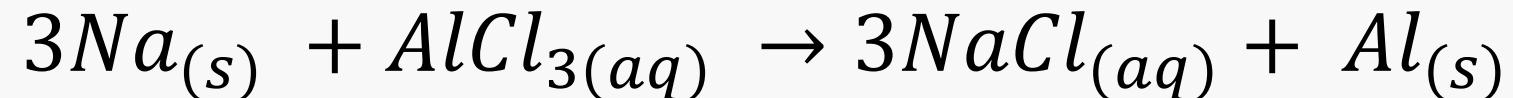
- ii. Write equation by using symbols and formulae.



- iii. Balance the equation. The total atoms of products should be equal to the total atoms of the reactants.



- iv. Include the state symbols.



Try this

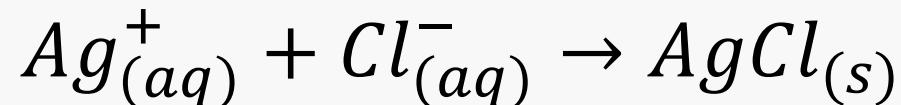
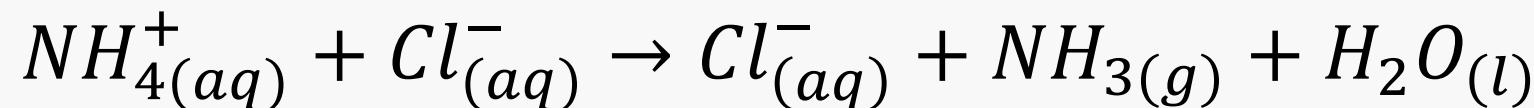
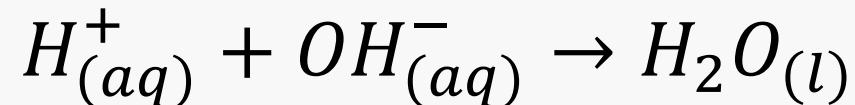
◦ Balance the following chemical equations

- i. $\text{Ca(OH)}_{2(s)} + \text{NH}_4\text{Cl}_{(aq)} \rightarrow \text{CaCl}_{2(aq)} + \text{NH}_{3(g)} + \text{H}_2\text{O}_{(l)}$
- ii. $\text{MgCO}_{3(s)} + \text{H}_2\text{O}_{(l)} + \text{CO}_{2(g)} \rightarrow \text{Mg}(\text{HCO}_3)_{2(aq)}$
- iii. $\text{CaCO}_{3(s)} + \text{HCl}_{(aq)} \rightarrow \text{CaCl}_{2(aq)} + \text{CO}_{2(g)} + \text{H}_2\text{O}_{(l)}$
- iv. $\text{H}_{2(g)} + \text{Br}_{2(g)} \rightleftharpoons \text{HBr}_{(g)}$
- v. $\text{NH}_4\text{Cl}_{(s)} \rightleftharpoons \text{NH}_{3(g)} + \text{HCl}_{(g)}$
- vi. $\text{NaNO}_{3(s)} \xrightarrow{\Delta} \text{NaNNO}_{2(s)} + \text{O}_{2(g)}$

iii) Ionic Equation

- Ionic equation is the equation which is written by means of ions.
- Ionic equation is the equation whereby spectator ions are omitted.
- **Spectator ions** are the ions which exist as reactants and products in a chemical equation.

Example:

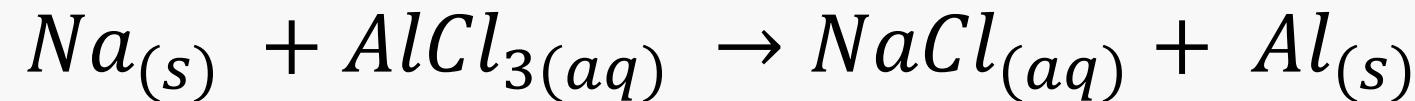


Steps to write ionic equation

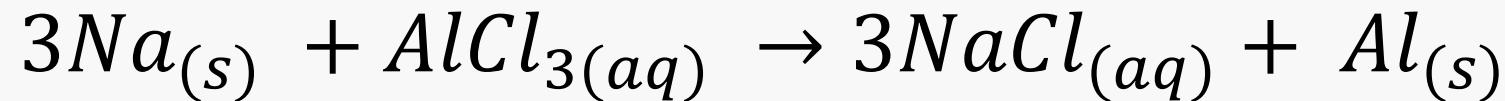
- i. Write equation in word

Sodium + Aluminium chloride → Sodium chloride + Aluminium

- ii. Write equation by using symbols and formulae.



- iii. Balance the equation



- iv. Split all soluble compounds into independent ion.



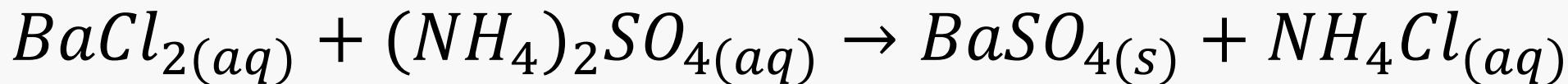
v. Cancel out spectator ions and write the net ionic equation.



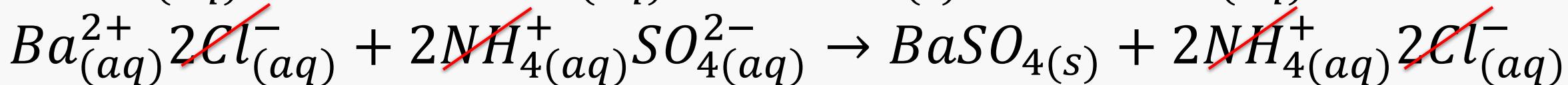
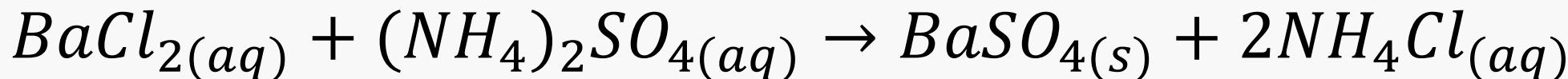
vi. Write the net ionic equation



Example: Write the ionic equation for the molecular equation below



Solution



❖ **Net ionic equation:** $\text{Ba}^{2+}_{(aq)} + \text{SO}_4^{2-}_{(aq)} \rightarrow \text{BaSO}_{4(s)}$

Try this

- Write the ionic equation for the following molecular equations
- $NaOH_{(aq)} + NH_4Cl_{(aq)} \rightarrow NaCl_{(aq)} + NH_3(g) + H_2O_{(l)}$
- $K_2CO_3_{(aq)} + HCl_{(aq)} \rightarrow KCl_{(aq)} + CO_2(g) + H_2O_{(l)}$
- $AgNO_3_{(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_3_{(aq)}$
- $Zn_{(s)} + HCl_{(aq)} \rightarrow ZnCl_2_{(aq)} + H_2(g)$

CHEMICAL REACTIONS

- A **chemical reaction** is a process that leads to the chemical transformation of one set of chemical substances to another.
- The chemicals which begin the reaction are called ***reactants*** and the chemical substances formed are called ***products***.

Types of chemical reactions

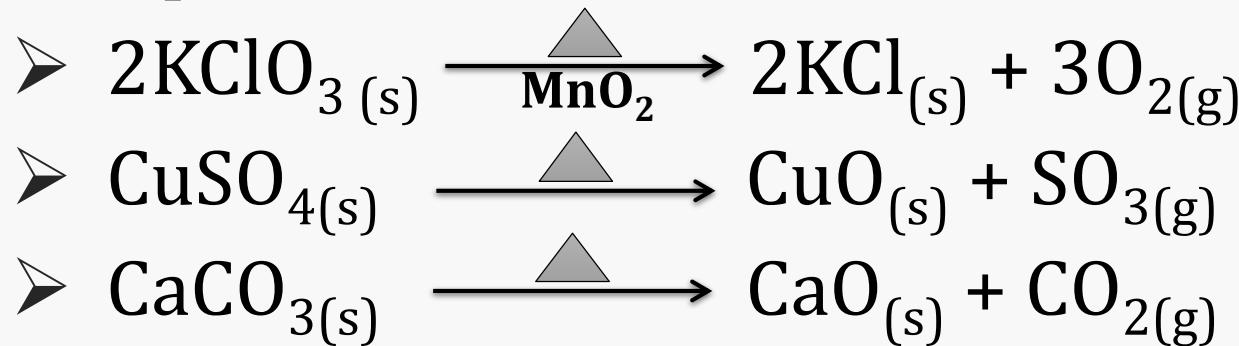
- There are four major types of chemical reactions, these are
 - i. Synthesis/combination reaction
 - ii. Decomposition reaction
 - iii. Single replacement reaction
 - iv. Double replacement reaction/double decomposition reaction

(i) Synthesis reaction (Direct Combination)

- Is the reaction where by two or more simple chemical species combine to form one more complex product.
- The general form of these reactions is: $A + B \rightarrow AB$
- Example:
 - $2\text{Mg}_{(\text{s})} + \text{O}_{2(\text{g})} \rightarrow 2\text{MgO}_{(\text{s})}$
 - $\text{C}_{(\text{s})} + \text{O}_{2(\text{g})} \rightarrow \text{CO}_{2(\text{s})}$
 - $\text{N}_{2(\text{g})} + 3\text{H}_{2(\text{g})} \rightleftharpoons 2\text{NH}_{3(\text{g})}$
 - $\text{Na}_2\text{O}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \rightarrow 2\text{NaOH}_{(\text{aq})}$

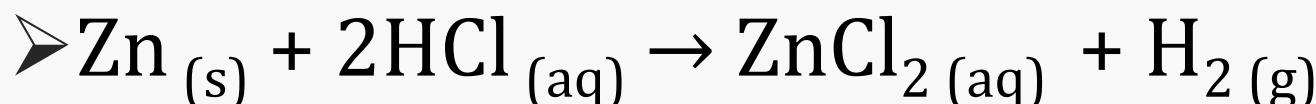
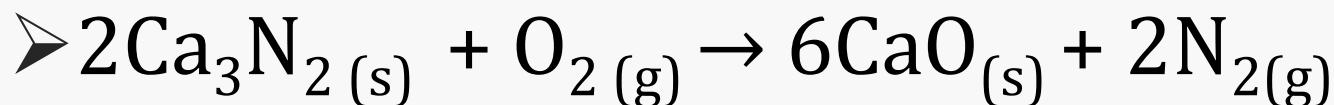
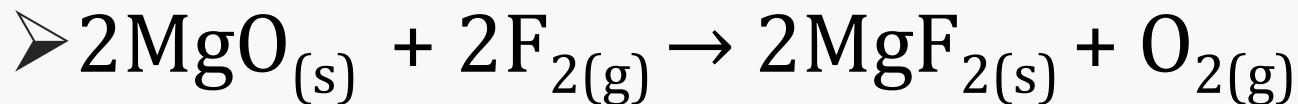
(ii) Decomposition reaction(Analysis Reaction)

- This is the reaction that involves breaking down of a compound into smaller chemical species.
- This can be categorized into three types which are ***thermal***, ***catalytic*** or ***electrical decomposition***.
- The general form of these reactions is: $AB \rightarrow A + B$
- It is thus the opposite of a synthesis reaction.
- Example:



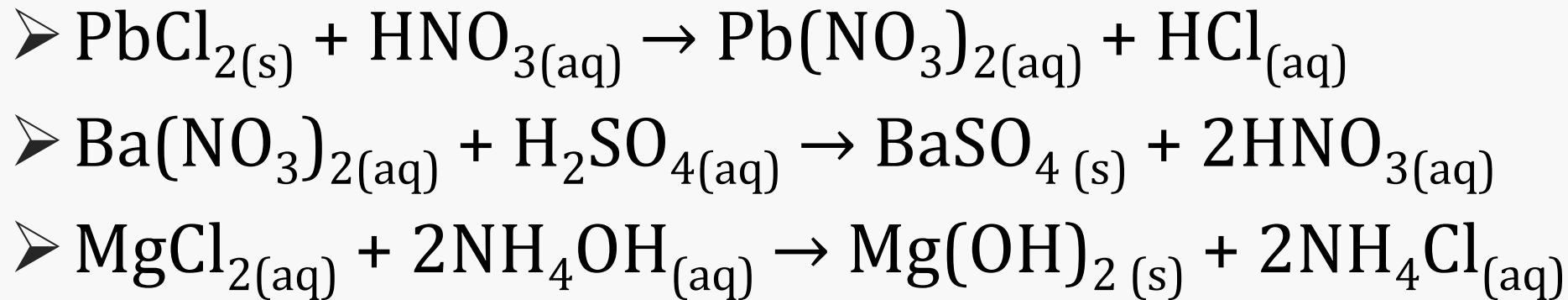
(iii) Single replacement(Substitution Reaction)

- Single replacement is characterized by one element being displaced from a compound by another element.
- Normally the less reactive element is displaced by more reactive one.
- The general form of these reactions is: A + BC → AC + B
- Example:



(iv) Double Displacement Reaction

- Is the reaction where by the anions and cations of two compounds switch places and form different compounds.
- It also known as ***double decomposition*** or ***Metathesis reaction***.
- The general form of these reactions is: AB + CD → AD + CB
- Example:



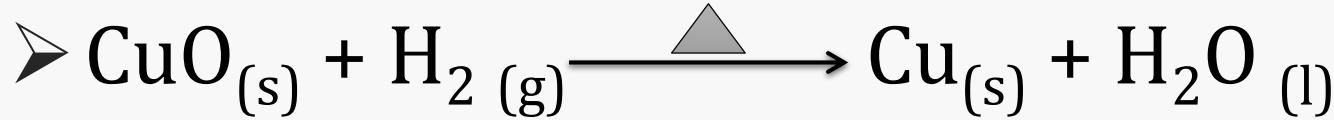
Other types of chemical reaction

- Other types of chemical reactions can be described as per the event occurring during chemical reaction.
- These may actually falls in the four types described earlier. These are;
 - i. Redox reaction
 - ii. Precipitation reaction
 - iii. Combustion reaction
 - iv. Neutralization reaction

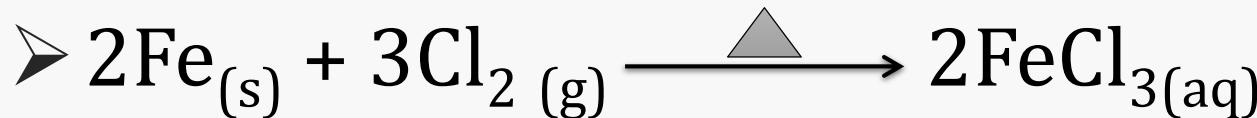
i) Redox reaction

- Is the reaction in which oxidation and reduction occurs at a time.
- This reaction may involve transfer of electrons from one species (reducing agent) to another (oxidizing agent).
- It may also involve exchange of oxygen or/and hydrogen.
- In practice, the transfer of electrons will always change the oxidation state.
- However there are many reactions that are classed as "redox" even though no electron transfer occurs (such as those involving *covalent* bonds)

◦ Example:



- ❖ Copper is reduced by losing oxygen hence it is an ***oxidizing agent***.
- ❖ Hydrogen is oxidized as it gains oxygen, hence it is a ***reducing agent***.



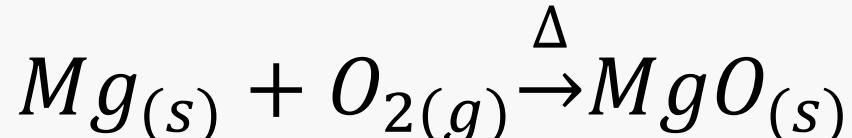
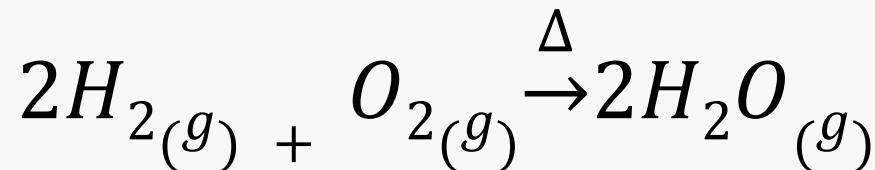
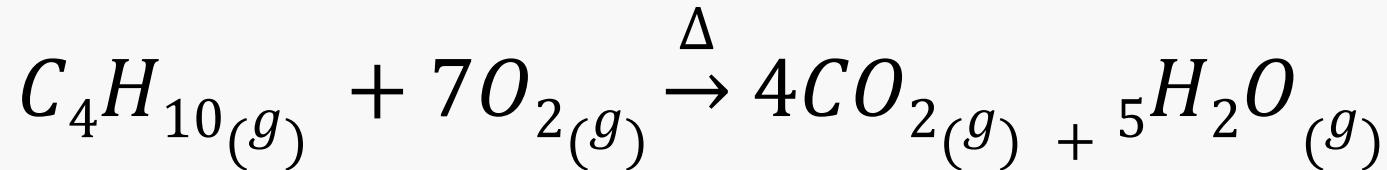
- ❖ Iron metal is oxidized by losing electron, hence it is a ***reducing agent***.
- ❖ Chlorine is reduced by gaining electron, hence it is an ***oxidizing agent***.

ii) Precipitation reaction

- Is the reaction in which two soluble compounds(aqueous) reacts to form soluble and insoluble compound.
- The insoluble compound appears as solid particles known as ***precipitates***, hence its name.
- It usually takes place when the concentration of dissolved ions exceeds the solubility limit and forms an insoluble salt.
- Example:
 - $\text{NaCl}_{(\text{aq})} + \text{AgNO}_{3(\text{aq})} \rightarrow \text{NaNO}_{3(\text{aq})} + \text{AgCl}_{(\text{s})}$
White precipitates
 - $\text{Ba}(\text{NO}_3)_{2(\text{aq})} + \text{H}_2\text{SO}_{4(\text{aq})} \rightarrow \text{BaSO}_{4(\text{s})} + 2\text{HNO}_{3(\text{aq})}$
**White
precipitates**

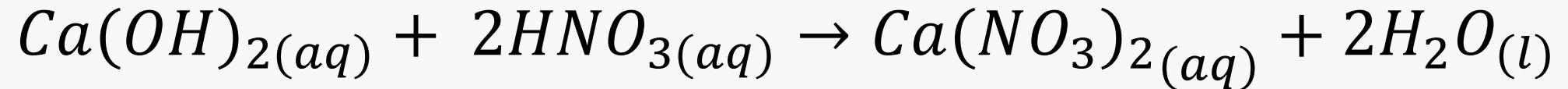
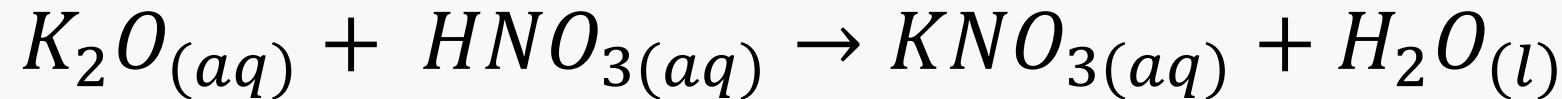
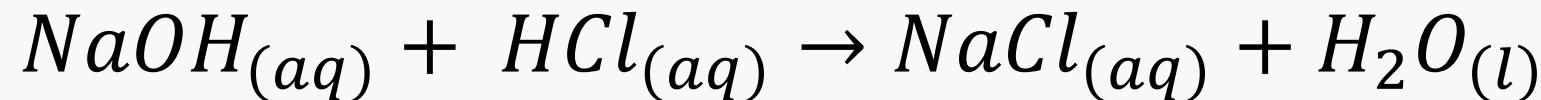
iii) Combustion reaction

- A combustion reaction is a type of redox reaction in which a combustible material combines with an oxidant to form oxidized products.
- This reaction is normally exothermic.
- The oxidant involved is usually atmospheric oxygen.
- Example:



iv) Neutralization reaction

- This is the reaction that occurs between an acid and a base to form salt and water only.
- It is a type of double displacement reaction.
- The H⁺ ion in the acid reacts with the OH⁻ ion in the base to form water and an ionic salt:



TOPIC 2

HARDNESS OF WATER

2.1 The Concept of Hardness of Water

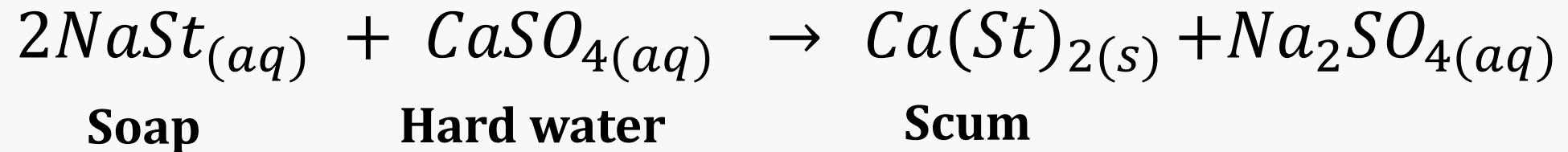
- As water percolates through deposits of *limestone*, *chalk* or *gypsum* which are largely made up of carbonates and bicarbonates, sulphates of calcium and magnesium it dissolves these salts and direct them together with water in rivers, lakes and oceans.
- Water containing these dissolved minerals is said to be hard water.
- Rain water forms lather easily with soap, thus it is said to be *soft water*.

- **Hard water** is the one which does not form lather easily with soap.
- **Hardness of water** is the measure of the amount calcium and magnesium salts in water.

Formation of scum

- Scum is a layer of debris or froth on the surface of a liquid. It is formed when hard water reacts with soap.
- Soap is made up of sodium and potassium salts of organic acids called **stearic acid**.
- This salt(scum) is simply represented as $Ca(St)_2$ or $Mg(St)_2$.

- When this salt from soap react with calcium or magnesium ions from hard water forms ***scum***(sodium stearate/potassium stearate)



- The problem of scum formation only occurs with soaps. Soapless detergents do not produce scum.

soft water

hard water



lather

SCUM

2.2 Causes of water hardness

- It caused by dissolved calcium and compounds, these includes
 - i) Calcium sulphate
 - ii) Calcium hydrogen carbonate
 - iii) Magnesium sulphate
 - iv) Magnesium hydrogen carbonate
- These salts dissolves in water when it passes through rocks which are largely made up of **calcium** and **magnesium** *carbonates, bicarbonates, chlorides and sulphates.*

Rocks which causes hardness of water

- The rocks which causes hardness of water are
 - i) Limestone rock($CaCO_3$)
 - ii) Gypsum rock($CaSO_4$)
 - iii) Chalk rock($CaCO_3$)
 - iv) Dolomite rock($MgCO_3$)

2.3 Types of Hardness of Water

- It classified into two types include
 - i) Temporary hard water
 - ii) Permanent hard water

2.3.1 Temporary hard water

- Temporary hard water is the one which can be softened by boiling.
- Temporary hard water is caused by dissolved hydrogen carbonates of calcium or magnesium.
- It is formed when rain water pass through the limestone or chalk rocks($CaCO_3$) in presence of carbon dioxide gas.



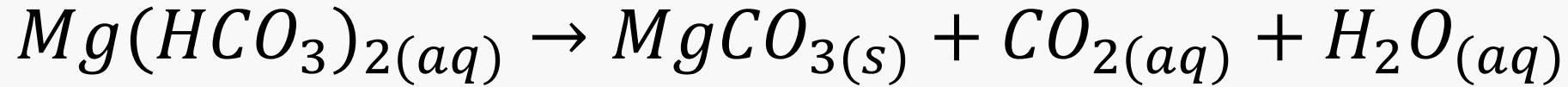
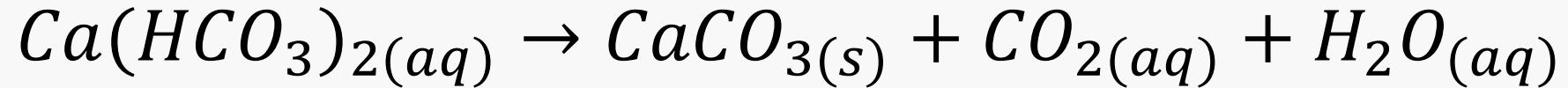
2.3.2 Permanent hard water

- Temporary hard water is the one which cannot be softened by boiling.
- Permanent hard water is caused by dissolved sulphates and chlorides of calcium or magnesium.
- It is formed when rain water pass through the gypsum ($CaSO_4$) rocks.

Methods of softening hard water

- i) Boiling
- ii) Distillation
- iii) Lime softening
- iv) Adding sodium carbonate (washing soda)
- v) Use of ion exchangers

i. ***Boiling:*** When Temporary hard water is boiled hydrogen carbonate decomposes to form precipitates of carbonate.

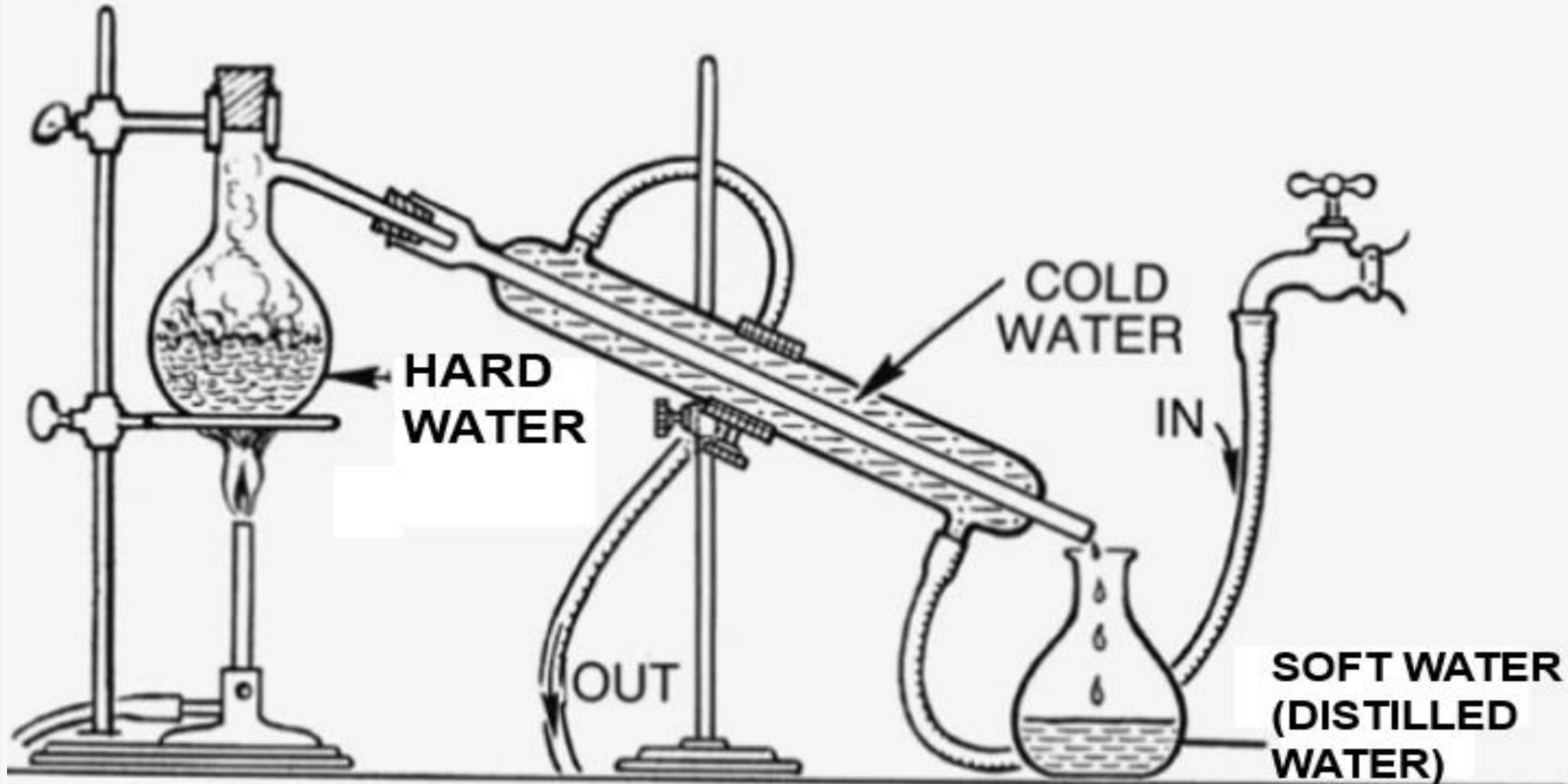


- Precipitate of carbonate filtered and leaves water soft.
- ❖ This method is only used to soften temporary hard water.

ii. ***Distillation:*** When hard water is heated it evaporates leaving the salts behind. The vapour is later cooled through the lie big condenser and soft water is obtained.

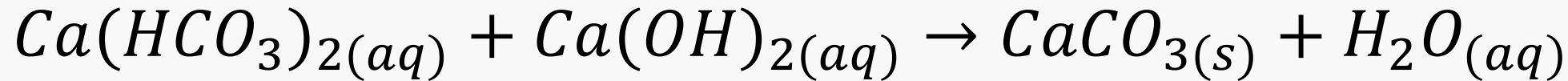
- This method can be used to soften both *temporary* and *permanent* hard water.

Distillation of hard water



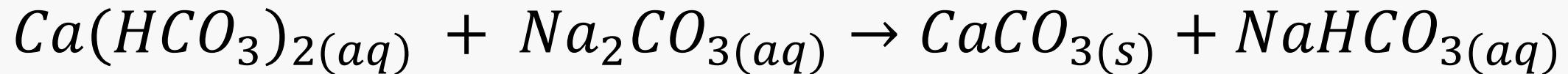
iii. Lime softening: This is done by adding lime water in water so as to soften it. This is used to soften temporary hard water.

- It has several advantages over the ion-exchange method, but is mainly suited to commercial treatment applications.



❖ This method is only used to soften temporary hard water.

iv. Adding Washing soda (sodium carbonate): When Sodium carbonate is added to precipitate calcium carbonate.

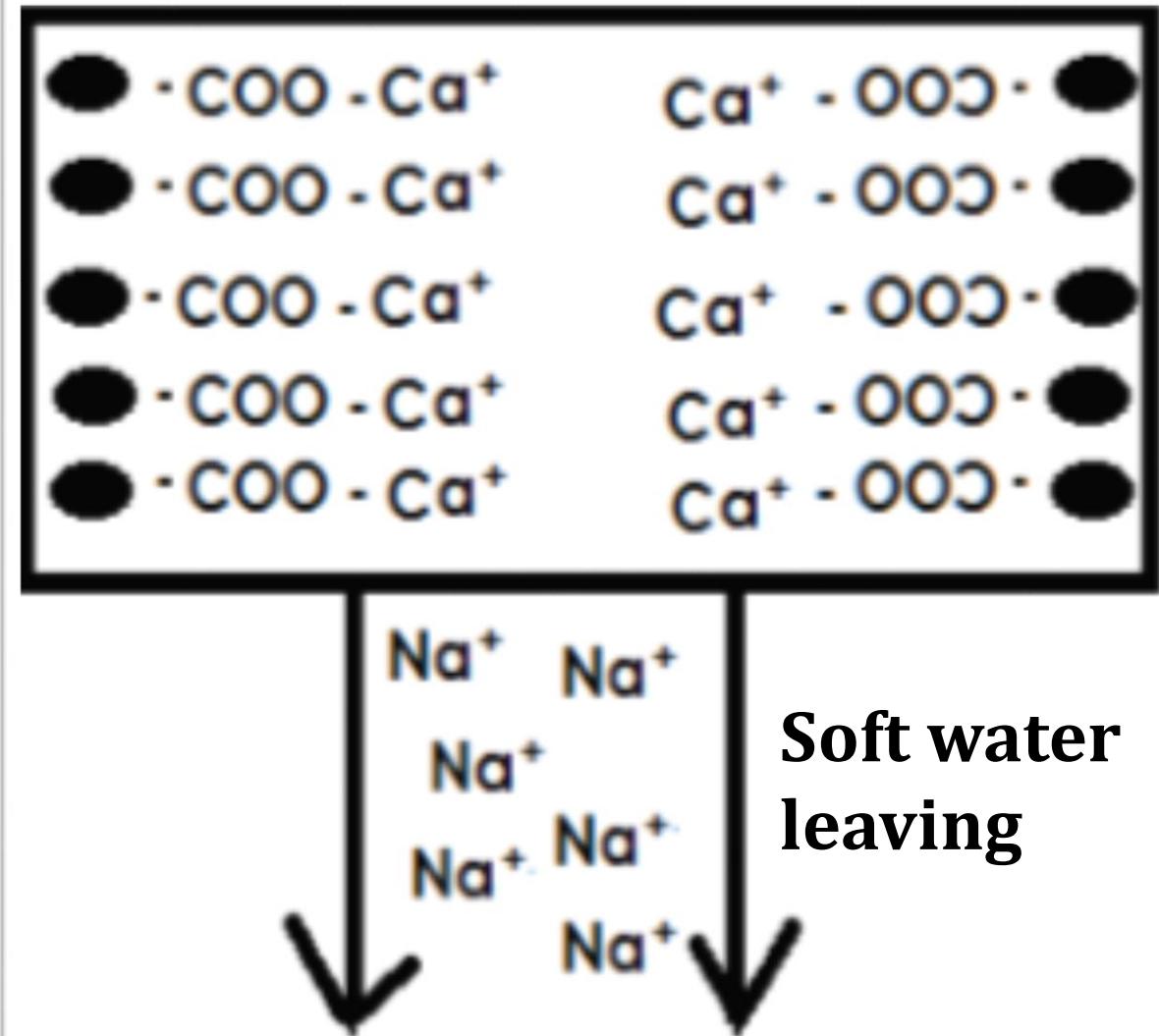
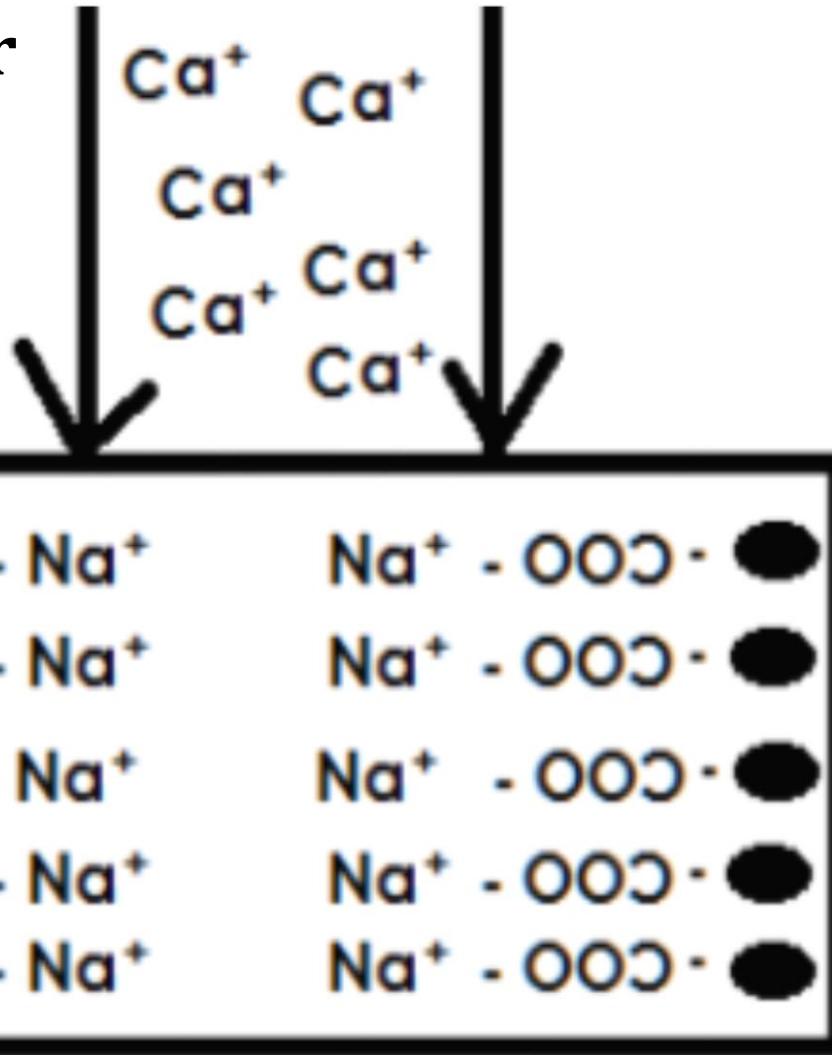


- This method can be used to soften both temporary and permanent hard water.

v. ***Use of ion exchangers:*** This is the containers full of small beads which consists special plastic called ion exchange resin.

- When tap water (temporary hard water) with calcium ions flow into ion exchanger, the calcium from hard water replace sodium ions in the ion exchanger to be used in the household as soft water.
- When all sodium ions replaced, ion exchanger cannot remove hardness any more.
- This is resolved by regeneration of the resin by pouring a concentrated solution of sodium chloride into exchanger.
 - ❖ This method can be used to soften both temporary and permanent hard water.

Hard water
entering



Advantage of hard water

- i) It tastes better due to dissolved compounds
- ii) It is better for brewing beer. Calcium ion promotes clarity, flavour and stability in the finished beer.
- iii) It provide useful calcium for growth of bones and teeth.
- iv) Formation of lime scale (insulation) in pipes which prevent them from rust.
- v) Toxic lead water pipes are less soluble in hard water.
- vi) Help in the formation of strong shells in some aquatic animals.

Disadvantage of hard water

- i) It need more soap than soft water.
- ii) Scum leaves marks on clothing and in baths.
- iii) It spoil the quality some special finishes on clothes.
- iv) It causes lime scale layer in hot water pipes, boilers, kettles.
This can make them less efficient and even block them.
- v) It's boiling point is elevated hence more fuel is used to boil it.



Hard Water

Lime scale



Soft Water



More questions

- 1.(a) Distinguish temporary hardness from permanent hardness of water.
- (b) With the help of chemical equations, explain how you can remove each type of water hardness in g(a).
2. With the aid of a chemical equation, briefly explain how
 - i) Temporary hardness of water can be removed by boiling.
 - ii) permanent hardness of water can be removed by chemical means.

TOIPC 3

ACIDS, BASES AND

SALTS

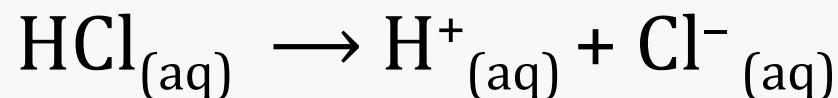
3.1 Acids and Base

3.1.1 Acids

- Acid is a chemical compound when dissolved in water produce hydrogen ions (H^+) as only positive charged ions.
- Word acid comes from the Latin word **acidus** which means **sour**.
- An acid can either be in solid form or gas.

Example:

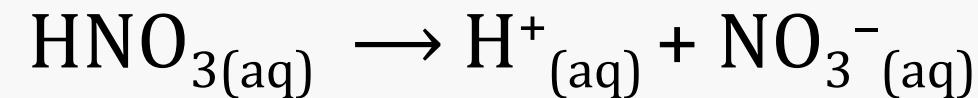
➤ Hydrochloric acid(HCl):



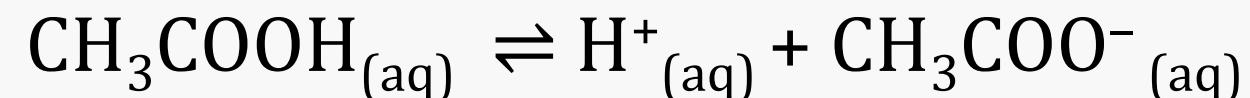
➤ Sulphuric acid (H_2SO_4):



➤ Nitric acid (HNO_3):



➤ Acetic acid (CH_3COOH):



Acids found in our daily life

S/N	ACID	NATURAL SOURCE
1.	Citric acid	Lemon juice, lime juice, orange juice, straw berry, grapes.
2.	Tartaric acid	Baking powder, health salt and pine apple
3.	Lactic acid	Sour milk
4.	Acetic acid	Vinegar
5.	Ascorbic acid	Orange juice
6.	Oxalic acid	Spinach
7.	Formic acid	Bee sting, nettle sting
8.	Tannic acid	Tea leaves
9.	Carbonic acid	Soft drinks

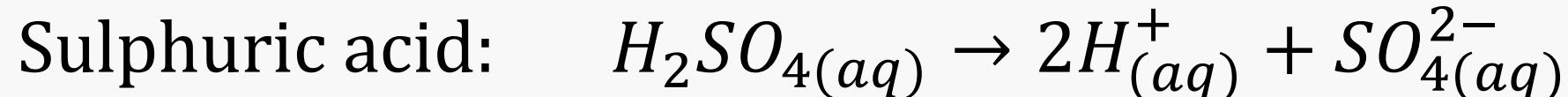
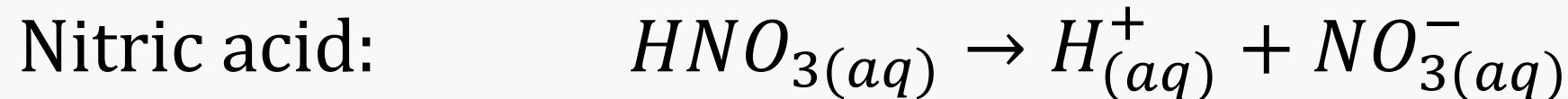
Types of acids

- Acids can be classified depending on two criteria
 - a) Basing on the strength of the acid.
 - b) Basing on the nature of the acid.

a) Basing on the strength of the acid.

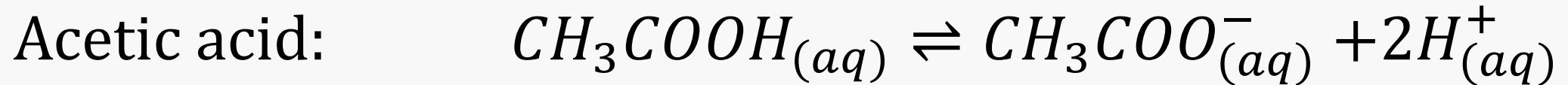
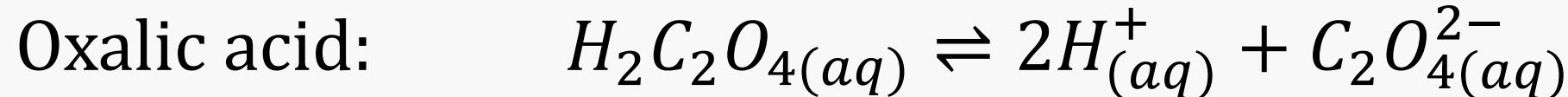
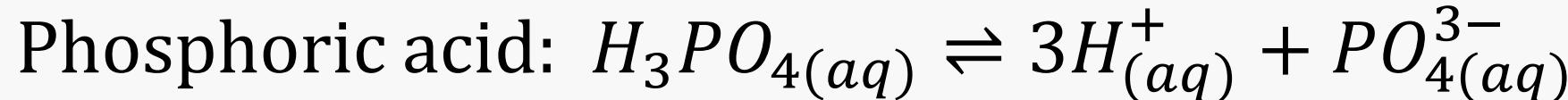
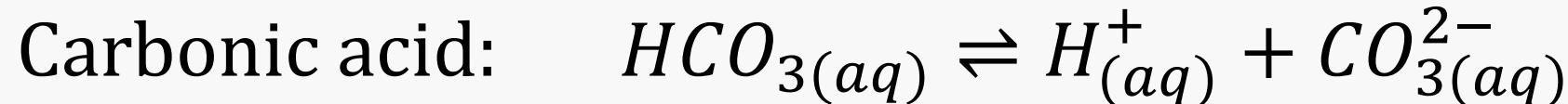
i) **Strong acids**: Are the acids which dissociate completely to form free ion in aqueous solution.

- These acids release a lot of hydrogen ions.



ii) Weak acids: Are the acids which dissociate partially to form free ion in aqueous solution.

- These acids release very few amount of hydrogen ions.



b) Basing on the nature of the acid

i) Organic acids are the acids that is derived from organic compounds.

- Most of the organic acids are found in our daily life.

Example: Acetic acid, ascorbic acid, tannic acid, oxalic acid.

ii) Inorganic acids are the acids which derived from inorganic compounds.

- They are also known as ***mineral acids***.

Example: nitric acid, phosphoric acid,

Basicity of an acidity

- *Basicity of an acid* is the number of ionizable hydrogen ions present in one molecule of an acid.
- According to basicity, acids can be grouped into two
 - i. *Monobasic (monoprotic) acids*: Are the acids which have one ionizable hydrogen ion.

Example: Hydrochloric acid: $HCl \rightarrow H^+ + Cl^-$

Hydrofluoric acid: $HF \rightarrow H^+ + F^-$

Nitric acid: $HNO_3 \rightarrow H^+ + NO_3^-$

ii. Polybasic (polyprotic) acids: Are the acids which have more than one ionizable hydrogen ions.

Example:

Dibasic acid: Sulphuric acid $H_2SO_4 \rightarrow H^+ + SO_4^{2-}$

Tribasic acid: Phosphoric acid $H_3PO_4 \rightarrow H^+ + PO_4^{3-}$

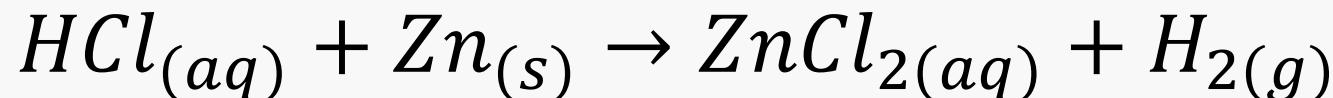
Dibasic acid: Carbonic acid $H_2CO_3 \rightarrow H^+ + CO_3^{2-}$

Physical properties of acids

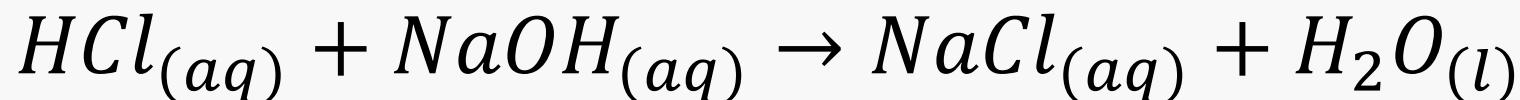
- i. Acid has a sour taste.
- ii. Acid turn blue litmus paper red.
- iii. Acid is corrosive.
- iv. Acid is electrolyte.

Chemical properties of acids

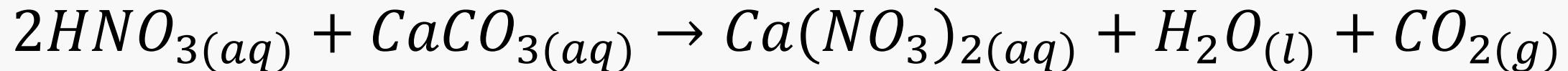
i. Acid React with metal to liberate hydrogen gas



ii. Acid React with base to form salt and water



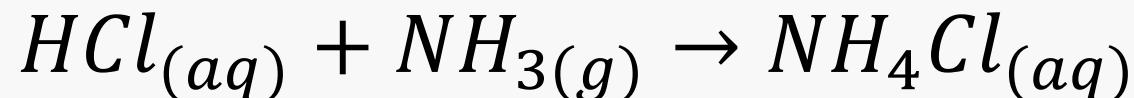
iii. Acid React with carbonates to form salt, carbon dioxide and water.



iv. Acid React with hydrogen carbonates to form salt, carbon dioxide and water.



v. Acid React with ammonia gas to form ammonium salts



3.1.2 Base

- Base is a chemical compound which reacts with hydrogen ions.
- Base is a chemical compound when dissolved in water produce hydroxyl ions OH^- or O^- as the only negative charged ions.
- **Examples:**

Sodium hydroxide (NaOH)

Potassium hydroxide (KOH)

Calcium oxide (CaO)

Lithium oxide (Li_2O)

Common names of the bases

S/N	Chemical name	Common name	Formula
1.	Potassium hydroxide	Caustic potash	KOH
2.	Sodium hydroxide	Caustic soda Soda lime	<i>NaOH</i>
3.	Calcium oxide	Quick lime	CaO
4.	Calcium carbonate	Limestone	<i>CaCO</i> ₃
5.	Sodium carbonate	Washing soda	<i>Na</i> ₂ <i>CO</i> ₃
6.	Ammonium hydroxide	Ammonia solution Aqueous ammonia	<i>NH</i> ₄ <i>OH</i>
7.	Calcium hydroxide	Lime water Slaked lime	<i>Ca(OH)</i> ₂

Alkalies

- Alkali is a basic oxide or hydroxide which is soluble in water.

Example:

- Sodium hydroxide
- Potassium hydroxide
- Calcium hydroxide
- Ammonium hydroxide
- Sodium oxide
- Potassium oxide

NOTE: Calcium hydroxide is the least soluble common alkali.

Types of base

- There are two types of base

Strong bases: Are the bases which dissociate completely to form free ion in aqueous solution.

- These acids release a lot of hydroxide or oxide ions.



ii) Weak acids: Are the acids which dissociate partially to form free ion in aqueous solution.

- These bases release very few amount of hydroxide or oxide ions.

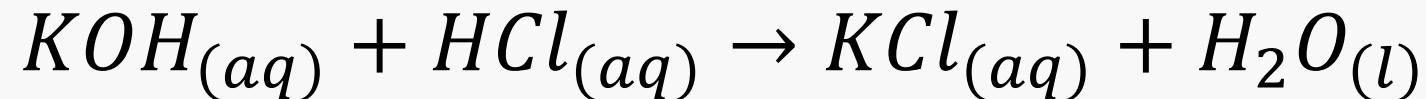


Physical properties of base

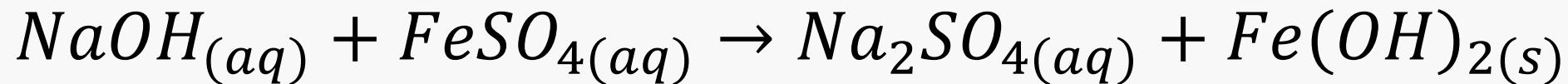
- i. Have a bitter taste.
- ii. Turn red litmus paper blue.
- iii. Have a soap or slippery feel.
- iv. Most insoluble in water
- v. They are corrosive.
- vi. They are electrolytes

Chemical properties of Base

- Base react with acid to form salt and water



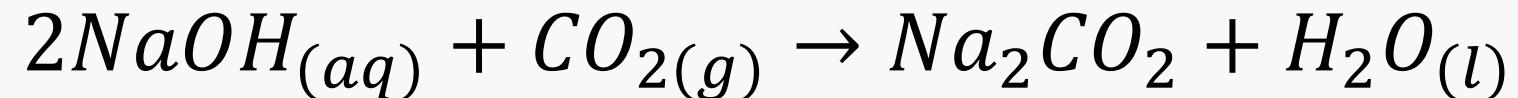
- Alkali precipitate insoluble metal hydroxides from their salt solution.



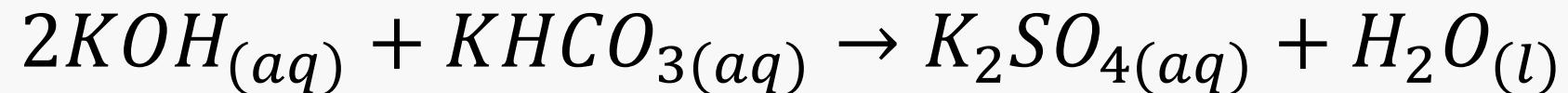
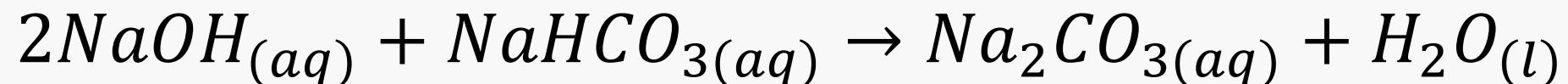
- Alkalies react with ammonium salt to produce ammonia gas, water and salt.



- Base react with carbon dioxide to form metal carbonate and water.



- Alkalies react with acidic salt to form normal salt.



3.2 Indicators

- Indicator is a chemical substance which helps in the visual detection of completion of titration.

OR

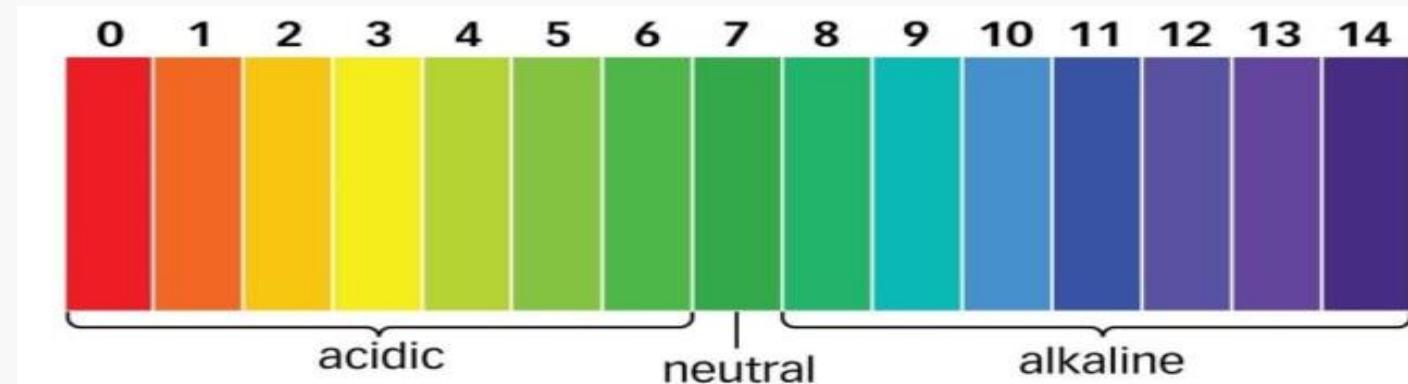
- Indicator is a chemical compound that shows a definite colour change in acid or base.

Common indicator used in the laboratory

Indicator	Colour	
	In base	In acid
Litmus	Blue	Red
Methyl orange(M.O)	Yellow	Pink
Phenolphthalein(P.O.P)	Pink	Colourless

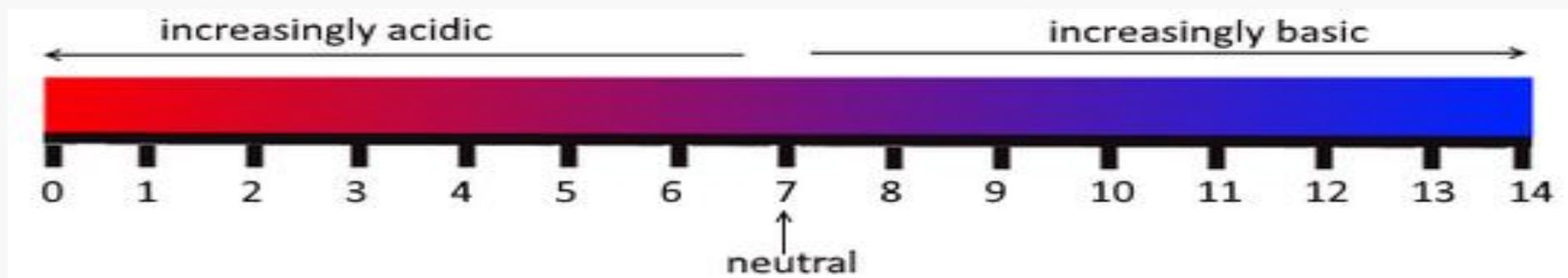
Universal indicator

- Universal indicator is a mixture of different simple indicator.
- It indicates pH values usually over a range of about 3 – 11 by successive changes in colour.
- Normally two drops of indicator solution are added to 10cm^3 of a test solution, with shaking.
- The colour developed is compared with a chart provided by a supplier to determine the pH of the solution.



pH scale

- pH scale is the numeric scale that indicates the degree of acidity and alkalinity of a solution.
- Specification of the pH values
 - pH=7 means neutral
 - pH<7 means acid
 - pH>7 means base



- **pH of a substance** is a measure of acidity or alkalinity of a water soluble substance.
- Mathematically:
- **pH of a substance** is the negative logarithm of hydrogen ions concentration of a solution.

$$pH = -\log[H^+]$$

Where: $[H^+]$ = Hydrogen ions concentration

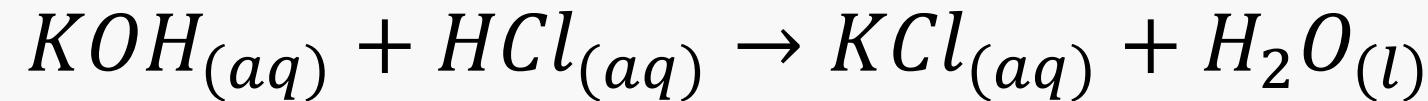
pH - metre

- **pH-metre** is an electronic device which is used to measure the pH of a substance.
- It consist of a glass electrode which is connected to the meter that measures and display the pH of a substance.



Neutralization

- Neutralization is the reaction between base and acid to form salt and water.



- Volume of acid neutralize alkalis is called Titre.

Application of neutralization

i. Treating insect stings and bites

- Acidic liquids injected into the skin by ant bite, bee sting and nettle sting is neutralized by rubbing with baking soda ($NaHCO_3$) in affected area(cucumber, ash and avocado may also be used).
- Wasp sting are alkaline and can be neutralized by vinegar (acetic acid).

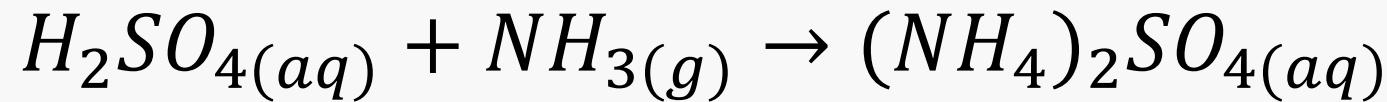
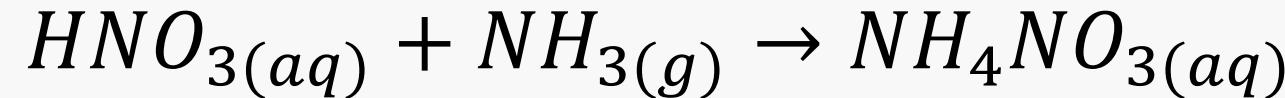
ii. Control of soil pH: Neutralization can be applied in amending the soil pH of the field.

- When soil is too acidic lime materials include quick lime (CaO), Slaked lime (CaOH) and Calcium carbonate (CaCO_3) added to neutralize soil.
- When soil is too much alkaline iron sulphate, ammonium based fertilizer and organic matter may be used.

iii. Relieving indigestion: Stomach pain (indigestion) or heartburn caused by presence of excess acid (HCl) in the stomach. It is relieved tables contain magnesium bicarbonates or sodium bicarbonates.

iii. Treating factory wastes: Waste from factories often contains acid which can cause death to living organism either land or water bodies. To prevent these lime materials added to neutralize wastes from industries.

iv. Manufacture of fertilizers: Ammonium fertilizer produced from neutralization between ammonia gas with mineral acid.



v. Detoxification of gaseous by-products: Acidic gases like carbon dioxide, sulphur dioxide and nitrogen dioxide produced by the industries causes acidic rainfall.

- Their effect is reduced by fitted scrubbers which gleans(extract) the gas from smoke.

3. Salts

- Salt is compound formed when all or part of hydrogen atom of an acid is replaced by metal or metallic radical(eg NH_4^+).
- Therefore a salt has basic part (metallic radical) which is positive ion and acidic part (non-metallic radical) which is negative ion.
- Example

Salt	Basic radical	Acidic radical
Ammonium sulphate- $(NH_4)_2SO_4$	NH_4^+	SO_4^{2-}
Sodium hydrogen carbonate- $NaHCO_3$	Na^+	HCO_3^-

Natural source of salt

- i. Lakes, seas and ocean water: These contains very high amount of dissolved chlorides of sodium and magnesium. This salt can easily be collected by evaporation.
- ii. Underground depositions: The underground deposition consist of mainly chlorides of sodium and potassium.
- iii. Example: Uvinza – Kigoma
- iv. Fresh foods: Some foods that are nutrient dense posses naturally occurring amount of salts especially some varieties of beans. Example: Mung beans and garbanzo beans

Garbanzo beans (chick pea)



Mung beans



Types of salts

◦ There are two major types of salts, these are;

i. ***Normal salt:*** This is the salt formed when all hydrogen of an acid is replaced by a metal or metallic radical.

Example: $NaCl$, NH_4Cl , K_2SO_4

ii. ***Acidic salt(bisalt):*** This is the salt formed when part of hydrogen of an acid is replaced by metal or metallic radical. The hydrogen that is left in a salt gives it's acidic nature.

◦ In naming these salts the word “**bi**” or “**acid**” can be used in the place of hydrogen.

◦ Example:

- $NaHCO_3$ – Sodium bicarbonate
- $KHSO_4$ – Potassium acid carbonate
- $Ca(HCO_3)_2$ – Calcium hydrogen carbonate

Other types of salts

i. **Basic salt:** This is the salt formed when amount of acid required to neutralize an alkali (base) is insufficient.

Example:

- Basic zinc chloride – $[ZnCl_2 \ Zn(OH)_2]$
- Basic lead(ii) carbonates - $[PbCO_3 \ Pb(OH)_2]$

ii. **Double salt:** This is the salt formed by intimate association of two simple salts when crystallize from their solution mixture.

Example:

- Potassium alum - $K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$
- Ammonium alum - $(NH_4)_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$

iii. Mixed salt: This is the salt that contains more than one acidic radical or basic radical.

Example:

- Potassium calcium phosphate - $KCaPO_4$
- Sodium potassium carbonate - $NaKCO_3$

iv. Complex salt: This is the salt that contain at least one complex entity.

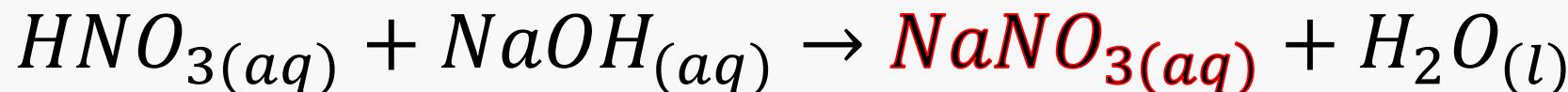
Example:

- Sodium aluminate
- Potassium ferrocynide

Preparation of salts

- ***Normal salts***

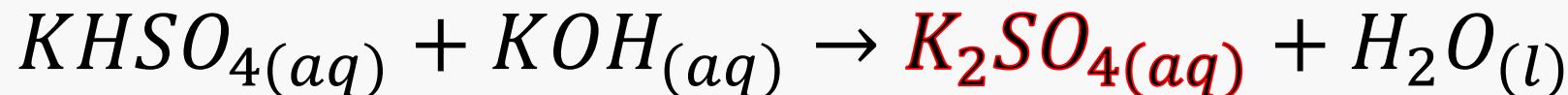
- i. Reaction between monobasic acid and base



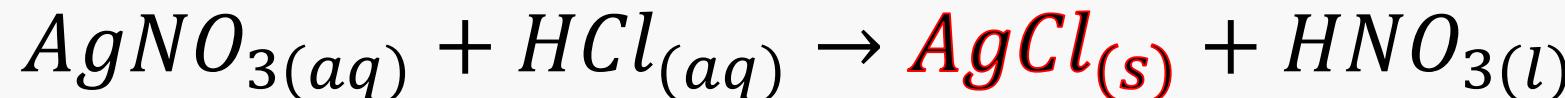
- ii. Reaction between metals and dilute acids



- iii. Reaction between acidic salt and base

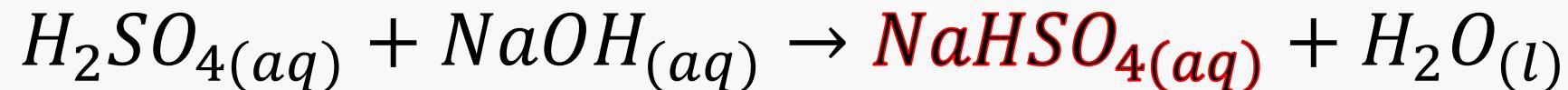


- iv. Precipitation reaction

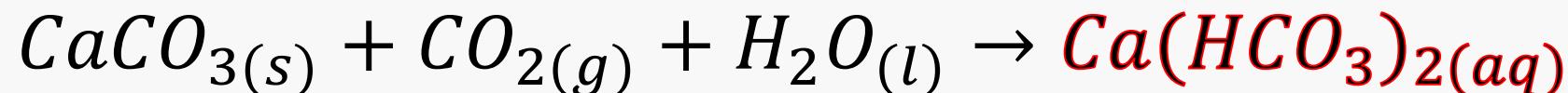


◦ ***Acidic salts***

i. Reaction between polybasic acid and base



ii. Reaction between metal carbonates and carbon dioxide in presence of water



iii. Reaction between metal carbonates and acids.



Properties of salts

- Will be covered in last topic (compound of metal)

Deliquescence and Efflorescence

- ***Deliquescence*** is the process by which a compound absorbs water from the atmosphere and form solution.
- A substance which absorbs water from the atmosphere and form solution is known as *deliquescent* (*verb: deliquesce*).

Example: $CaCl_2$, $NaOH$, KOH , P_4O_{10} (*in water it form H_3PO_4*)

- ***Hygroscopic*** is a substance which attracts water from the atmosphere.

Example: CaO , KCl , $NaCl$, *conc* H_2SO_4 , $CuSO_4$

- ***Efflorescence***: is the process by which a compound loses it's water to the atmosphere.
- A substance which loses water to the atmosphere and is known as *efflorescent* (*verb: Effloresce*).

Example: $CuSO_4 \cdot 5H_2O$, $FeSO_4 \cdot 7H_2O$

Uses of salts

- Control of soil pH: Calcium oxide added when soil is acidic and calcium sulphate (gypsum) added when soil is alkaline.
- Used as antacid: Magnesium sulphate (Epsom salt) and sodium hydrogen carbonate are used to relieve heartburn and acid in the stomach.
- Used as inorganic fertilizer: Inorganic fertilizers like ammonium sulphate, ammonium nitrate and calcium phosphate are salts.
- Used for food flavouring and preservation: Sodium chloride is used for seasoning, preserving food and adding flavour in food.

- Used for softening hard water: Sodium carbonates used to soften both permanent and temporary hard water.
- Used as drying agent (desiccant): Some salts which are deliquescent like Calcium chloride can be used as a drying agent.
- Copper (II) sulphate is used as a fungicide.

TOPIC 4

MOLE CONCEPT AND RELATED CALCULATIONS

4.1 The Mole as a Unit of Measurement

- Mole of substance is the amount of substance containing as many elementary units as there are atoms in 12 g of carbon-12 isotope.
- A mole is an SI unit of amount of substance. It's short form is "***mol***"
- The elementary particles present in a mole of a substance are ions, molecules or electrons.

Mole and other units

Unit	Number of objects per unit
Pair	2 objects. Example: shoes, socks
Dozen	12 Objects. Example: Cups, plates
Gross	144 Objects. Example: Chalks in the box
Ream	500 Objects. Example: sheets of papers
Mole	6.02×10^{23} Particles. Example: atoms, ions, molecules <ul style="list-style-type: none">• 1 mole of atoms = 6.02×10^{23} atoms• 1 mole of ions = 6.02×10^{23} ions• 1 mole of molecules = 6.02×10^{23} molecules

- The number of particles present in one mole of a substance is constant for any chemical substance and it is called ***Avogadro's constant***. It is denoted by “L” or “ N_A ”
- Therefore, a mole is the Avogadro's number of atoms, ions or molecules of a substance.

4.2 Application of the Mole Concept

- **4.2.1 Molar mass(M_m)**
- ***Molar Mass*** is the mass of one mole of the substance expressed in grams per mole. SI Unit of molar mass is g/mol.
- The molar mass of a substance is the same as relative molecular mass of a compound in magnitude. But the relative molecular mass of a substance has no SI unit.
- Example:
 - CO_2** – It's molar mass is 44g/mol, but it's relative molecular mass is 44.
 - K** – It's molar mass is 39g/mol, but it's relative atomic mass is 39.

Difference between molar mass and relative molecular mass

S/N	Molar Mass	Relative molecular mass
1.	It has SI units (g/mol)	Has no SI units
2.	It is the mass of one mole of a substance.	It the sum of atomic masses of all the atoms in a compound.
3.	The term molar mass is used in compounds and elements.	The term relative molecular mass is applied in compounds or molecule.

- In expressing the amount of substance (moles) one should specify the exactly quantity.
- Example:
- N_2 – 1 mole of nitrogen molecule containing two atoms.
- Mathematically, the amount of substance (n) is the ratio of mass(M) over molar mass (M_m).

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

Example 1

- The relative atomic mass of magnesium is 24. How many moles are there in 35.8 g of magnesium?

Data given:

- Mass, M = 35.8 g
- Molar mass, M_m = 24

◦ From

$$n = \frac{M}{M_m}$$
$$n = \frac{35.8}{24}$$

$$n = 1.49\text{mol}$$

Amount of substance, n is
1.49mol

Example 2

- Calculate the mass of 2 moles of sodium chloride. (Given Na=23, Cl=35.5)

Data given

- Amount of NaCl, n = 2mol
- Molar mass of NaCl, M_m =
- Required: Mass of NaCl, M

- $M_m = 23 + 35.5 = 58.5 \text{ g/mol}$
- Then;

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

$$M = n \times M_m$$

$$M = 2 \text{ mol} \times 58.5$$

$$M = 117 \text{ g}$$

Mass of substance 117g

4.1.2 Molar Volume

- Molar volume is the volume of one mole of gas at standard temperature and pressure.
- It is denoted by GMV or V_m . Its SI unit is dm^3 or cm^3 .
- The particles which make up gases are the molecules.
- One molecule of a gas contains 6.02×10^{23} molecules.
- Avogadro's law states that, "Equal volume of gases at standard temperature and pressure contains the same number of molecules"
- The volume of one mole of a gas at S.T.P is $22.4dm^3$ or $22400cm^3$.

$$1dm^3 = 1L = 1000cm^3$$

- For any gas at S.T.P the amount of substance can be obtained as follow;

$$n = \frac{V}{V_m} \quad OR \quad n = \frac{V}{G.M.V}$$

Example 1

- Calculate the amount of substance present in 48.8L of oxygen at S.T.P.(G.M.V=22.4dm³)

Data given:

- Volume , V= 48.8L
- Molar volume, V_m = 22.4dm³
- Required: Amount of O_2 ,n

◦ From: $n = \frac{V}{V_m}$

$$n = \frac{48.8L}{22.4dm^3}$$

$$n = 2.18mol$$

Amount of O_2 is 2.18mol.

Example 2

- Determine the volume of 0.5 mol of carbon dioxide gas at S.T.P. (G.M.V=22.4L)

◦ Data given:

- Amount of CO_2 , $n = 0.5\text{mol}$
- Molar volume, $V_m = 22.4L$
- Required: Volume , V

- From the formula:

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

$$V = n \times V_m$$

$$V = 0.5\text{mol} \times 22.4L/\text{mol}$$

$$V = 11.2L$$

- Volume of CO_2 is 11.2L

Example 3

- Determine the volume of 5g of ammonia gas at S.T.P.
(G.M.V=22.4dm³, N=14, H=1)

Data given:

- Mass, M = 5g
- Molar volume, $V_m = 22.4\text{dm}^3$
- Molar mass, $M_m = ?$
- Required: Volume of NH_3

- $M_m = 14 + 1 \times 3 = 17$

- From: $n = \frac{M}{M_m}$

$$n = \frac{5\text{g}}{17}$$

$$n = 0.294\text{mol}$$

- Then, $n = \frac{V}{V_m}$

$$V = n \times V_m$$

$$V = 0.294\text{mol} \times 22.4\text{dm}^3/\text{mol}$$

$$V = 6.59\text{dm}^3$$

Volume of NH_3 is 6.59dm^3

Try this

1. Calculate the volume of
 - a) 0.2mol of chlorine gas
 - b) 5.6g of nitrogen gas
2. Determine the amount of substance in the following
 - a) 1.6g of magnesium oxide
 - b) 89.6ml of fluorine gas at S.T.P.

4.2.3 Change masses of solids or volumes of known gases to actual number of particles

- The number of particles in “n” moles of a substance is given by as follows:

$$N = n \times L$$

Where:

N = Number of particles

N = Amount of substance (in moles)

L = Avogadro's constant (6.02×10^{23} particles)

Example 1

- How many molecules are there in 20 litres of oxygen gas?
- Data given:
- Volume of O_2 , $V = 20L$
- Molar volume, $V_m = 22.4L$
- Avogadro's constant, $L = 6.02 \times 10^{23}$ molecules

$$n = \frac{V}{V_m}$$
$$n = \frac{20L}{22.4L/mol}$$
$$n = 0.893mol$$

Then, $N = n \times L$
 $N = 0.893mol \times 6.02 \times 10^{23}$
 $N = 5.38 \times 10^{23}$

Number of molecules is 5.38×10^{23}

Example 2

- Calculate the number of atoms present in 10g of carbon.

$$n = \frac{M}{M_m}$$
$$n = \frac{10g}{12g/mol}$$
$$n = 0.833mol$$

Data given

- Mass of C, M = 10g
- Molar mass of C, $M_m = 12g/mol$

Then, $N = n \times L$

$$N = 0.833mol \times 6.02 \times 10^{23}$$
$$N = 5.0 \times 10^{23}$$

Number of molecules is 5.0×10^{23}

Example 2

- Calculate the number of atoms present in 10g of carbon.

$$n = \frac{M}{M_m}$$
$$n = \frac{10g}{12g/mol}$$
$$n = 0.833mol$$

Data given

- Mass of C, M = 10g
- Molar mass of C, $M_m = 12g/mol$

Then, $N = n \times L$

$$N = 0.833mol \times 6.02 \times 10^{23}$$

$$N = 5.0 \times 10^{23}$$

Number of molecules is 5.0×10^{23}

Try this

1. Calculate the number of particles present in
 - a) 32g of water?
 - b) 3mol of chlorine gas?
 - c) 4.48 litres of Sulphur dioxide gas?
2. Calculate the volume of 7.0×10^{22} molecules of ammonia gas at s.t.p.

Number of particles in salts

- The particles which makes up salts are the ions.
- Therefore, in finding the number of particles(ions) in salts one should firstly find the amount of ions. This can be done by considering dissociation of the salt.
- For example, Magnesium chloride dissociates as follows:



1 mol 1mol 2mol

- This means;
- 1mol of magnesium chloride produces 3 moles of ions

Example 1

- Calculate the number of ions present in 4.26g of sodium sulphate.

Data given

- Mass of Na_2SO_4 , M = 10g

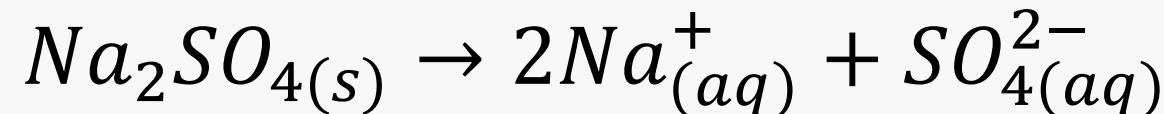
Molar mass of Na_2SO_4 , M_m

$$M_m = [(23 \times 2) + 32 + (16 \times 4)] g/mol$$

$$M_m = 142 g/mol$$

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

$$n = \frac{4.26g}{142g/mol}$$
$$n = 0.03mol$$



1 mole 2moles 1mole

1mol of Na_2SO_4 $\xrightarrow{\text{produce}}$ *3mol of ions*

0.03mol of Na_2SO_4 $\xrightarrow{\text{produce}}$ *xmol of ions*

$$n = \frac{0.03 \times 3}{1}$$

$$n = 0.09\text{mol}$$

Amount of ions is 0.09mol

Then, the number of ions is given by;

$$N = n \times L$$

$$N = 0.03\text{mol} \times 6.02 \times 10^{23}$$

$$N = 5.4 \times 10^{24} \text{ ions}$$

❖ Number of ions present in Na_2SO_4 is 5.4×10^{24} ions

Try this

- Find the number of ions in the following salts
 - i. 10g of calcium chloride.
 - ii. 2mol of ammonium sulphate.

4.2.4 Mole in stoichiometric Calculations

- **Stoichiometry** is the quantitative relationship between the reactant and the product.
- This shows how much of a substance reacts to form another substance.
- Example; $3H_{2(g)} + N_{2(g)} \rightleftharpoons 2NH_{3(g)}$
- *This can be read as;*
- ❖ 3moles of hydrogen gas reacts with 1mole of nitrogen gas to form 2moles of ammonia gas.
- The numbers 3,1 and 2 are called stoichiometric coefficients/values
- They tell us the proportions of reactants that form products.

- In any chemical reaction for the reaction to stop at least one of the reactant should have been finished.
- The reagent which is fully consumed is called limiting reagent or limiting reactant
- **Limiting reagent** is the reactant in the chemical reaction that determines the amount of product that is formed.

OR

- **Limiting reagent** is the reactant that is fully consumed when a reaction goes to completion.

Example 1

- 0.2mol hydrogen reacted with nitrogen to produce 2.24 L of ammonia at STP. Calculate
 - a) Amount of nitrogen reacted
 - b) Volume of nitrogen reacted
 - c) Find the amount of excess hydrogen.
- **Data given:**
- Volume of NH_3 , $V = 2.24L$
- Molar volume, $V_m = 22.4L$

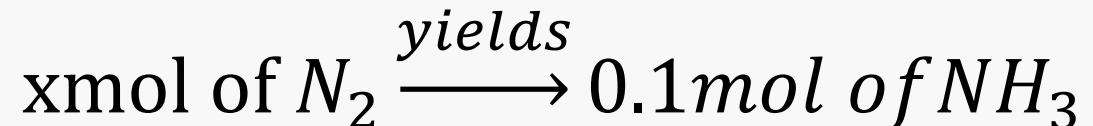
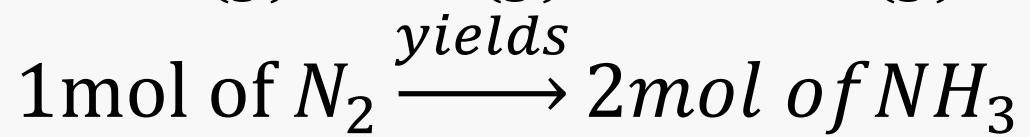
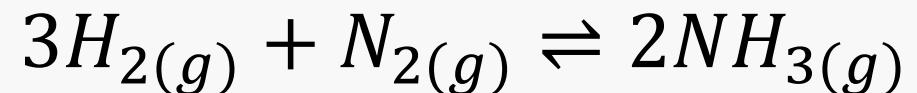
a) Required: Amount of N_2

Firstly find amount of NH_3

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

$$n = \frac{2.24L}{22.4L/mol}$$

Amount of NH_3 , $n = 0.1mol$



$$x = \frac{1 \times 0.1}{2}$$

$$x = 0.05\text{mol}$$

Amount of nitrogen is 0.05mol

b) Volume of N_2

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

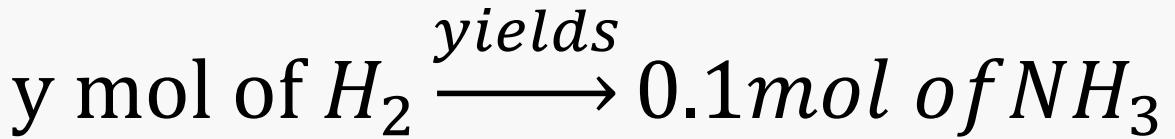
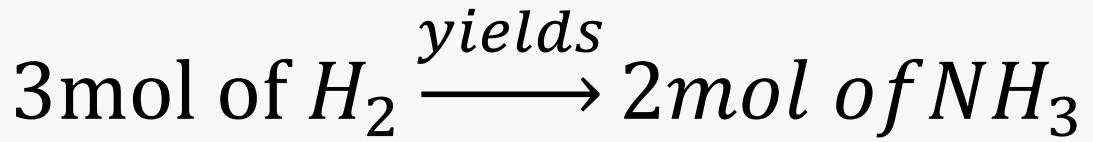
$$V = n \times V_m$$

$$V = 0.05\text{mol} \times 22.4\text{L/mol}$$

$$V = 1.12\text{L}$$

Volume of N_2 is 1.12L

c) Required: Amount of excess hydrogen



$$y = \frac{3 \times 0.1}{2}$$

$$y = 0.15\text{mol}$$

Amount of H_2 consumed is 0.15mol

Excess = 0.2mol - 0.15mol

Excess H_2 is 0.05mol

Try this

- When 0.12g of a metal was dissolved in excess dilute sulphuric acid, 112cm^3 of dry hydrogen at stp were evolved. Calculate the mass of the metal which replaces 1g of hydrogen, if the density of hydrogen at stp is 0.09gdm^{-3} . Identify the metal.

TOPIC 5

VOLUMETRIC ANALYSIS

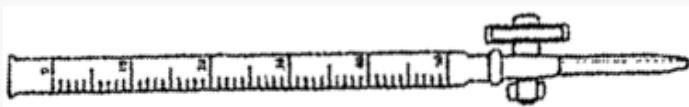
5.1 Standard Volumetric Apparatus

5.1.1 The concept of volumetric analysis

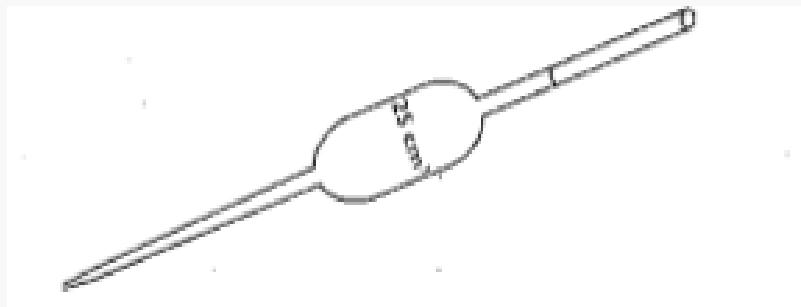
- Volumetric analysis is a chemical procedure for determining the concentration of a solution.
- It can also be termed as quantitative analysis because it deals with the determination of the amount of one constituent of a sample.
- The solutions concerned at this level include acid and alkali.

5.1.2 Volumetric Apparatus

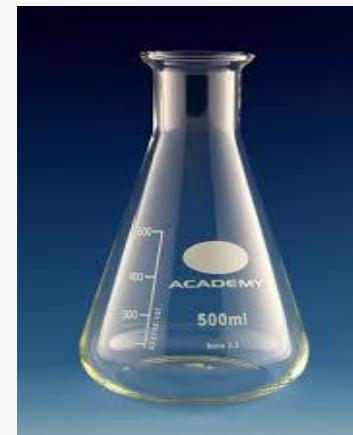
i. Burette



ii. Pipette



iii. Conical flask(Titration flask)



iv. Retort stand with clamp



v. Beakers



vi. Dropping bottle



vii. Measuring cylinders



viii. Wash bottle



ix. Dropper



x. Volumetric flasks



xi. Filter funnel



xii. White tile or paper



Mandatory Individual Assignment

- Explain the function of each apparatus used in volumetric analysis.

Important terms used

- i. **Standards solution** is solution of known concentration.
- ii. **Molar solution** is the solution which contains one mole of a solute in each litre of a solution.
- iii. **Mass concentration**(or simply concentration) is the mass of a substance in grams dissolved in 1litre of a solution. It's SI unit is g/L or g/dm^3 .
- iv. **Molar concentration**(molarity) is the amount of substance in moles dissolved in 1litre of a solution. It's SI unit is ***molar, M*** (mol/L or mol/dm^3).
- v. **Normal concentration**(normality) is the number of equivalents of a solute per litre of a solution. It's SI unit is ***normal, N (equiv/L*** or $equiv/dm^3$).

vi. *Equivalent of a solute* is the mass of a solute in gram per equivalent weight.

vii. *Equivalent weight (gram equivalent)* is the mass of a given substance which will combine with or displace a fixed quantity of another substance.

a) *Equivalent weight of an acid* is the weight of an acid that supplies one mole of protons ($1\text{mol} = 6.02 \times 10^{23} \text{protons}$).

b) *Equivalent weight of a base* is the weight of a base that reacts with one mole of protons.

Calculations involving mass Concentration and Molar concentration

- Formula for finding mass concentration

$$\text{Mass concentration, } C = \frac{\text{Mass of a solute, } M_s}{\text{Volume os a solution, } V_{sn}}$$

- Formulas for finding molar concentration(Molarity)

$$\text{Molar concentration, } M = \frac{\text{Mass concentration, } C}{\text{Molar mass, } M_m}$$

$$\text{Molar concentration, } M = \frac{\text{Mass of a solute, } M_s}{\text{Volume os a solution, } V_{sn} \times \text{Molar mass, } M_m}$$

$$\text{Molar concentration, } M = \frac{\text{Amount of solute, } n_s}{\text{Volume os a solution, } V_{sn}}$$

Examples

1. Find the molarity of a solution made by dissolving 4.24g of sodium carbonate solution in water to make solution.
2. A solution of hydrochloric acid was made by dissolving 5moles of hydrogen chloride gas in water to make 1500ml of a solution. Calculate it's concentration in;
 - a) Mole per litre,
 - b) Gram per litre.

Try this

1. How many number of moles present in the following;
 - a) 120 cm^3 of 0.6M Na_2CO_3 solution
 - b) 200 cm^3 of 0.1M HCl solution.
2. Calculate the mass of magnesium sulphate crystals required to make 250ml of 0.2M solution.

Calculations involving Equivalents and Normality

- Formula for finding equivalent of a solute

$$\text{Equivalent of a solute of a solute, } \text{Equiv}_s = \frac{\text{Mass of a solute, } M_s}{\text{Equivalent weight, } E_w}$$

- Formulas for finding normal concentration(normality)

$$\text{Normality, } N = \frac{\text{Mass concentration, } C}{\text{Equivalent weight, } E_w}$$

$$\text{Normality, } N = \frac{\text{Mass of a solute, } M_s}{\text{Volume os a solution, } V_{sn} \times \text{Equivalent weight, } E_w}$$

$$\text{Normality, } N = \frac{\text{Equivalent of a solute of a solute, } \text{Equiv}_s}{\text{Volume os a solution, } V_{sn}}$$

Examples

1) What is the equivalent weight of sulphuric (H_2SO_4). (assume all available protons are donated when acid react with base).

Solution

$$\text{Molar mass of } H_2SO_4 = (1 \times 2) + 32 + (16 \times 4)$$

$$= 98\text{g/mol}$$



The weight of one mole of H_2SO_4 is 98g.

- 1mole (98g) of H_2SO_4 can supply 2moles of protons to a base.

Therefore 98g of $H_2SO_4 \rightarrow 2\text{moles of protons}$
 y g of $H_2SO_4 \rightarrow 1\text{moles of protons}$

$$\text{Equivalent weight, } E_w \text{ of } H_2SO_4 = \frac{98g \times 1\text{mol}}{2\text{mol}}$$
$$E_w = 49\text{g/equiv}$$

❖ Equivalent weight of H_2SO_4 is 49g/equiv.

Note: g/equiv is the equivalent weight in grams.

2) A solution was made by dissolving 49g of phosphoric acid in enough water to make 600ml of solution. Assume that all available protons are donated when the acid solution reacts with a base, calculate;

- a) Equivalent weight of the acid.
- b) Equivalent of a solute.
- c) Normality of the acid.

Solution

Mass of a solute, $M_s = 49\text{g}$

Volume of solution, $V_{sn} = 600\text{ml}$

Molar mass of $H_3PO_4 = (1 \times 3) + 31 + (16 \times 4)$

$$= 98\text{g/mol}$$



The weight of one mole of H_3PO_4 is 98g.



◦ 1mole (98g) of H_3PO_4 can supply 3moles of protons to a base.

Therefore 98g of $H_3PO_4 \rightarrow 3\text{moles of protons}$

y g of $H_3PO_4 \rightarrow 1\text{moles of protons}$

$$\text{Equivalent weight, } E_w \text{ of } H_3PO_4 = \frac{98g \times 1\text{mol}}{3\text{mol}}$$

$$E_w = 32.7\text{g/quiv}$$

❖ Equivalent weight of an acid is 32.7 g/quiv

$$\text{Equivalent of a solute of a solute, } \text{Equiv}_s = \frac{49\text{g}}{32.7\text{g/quiv}}$$

$$\text{Equiv}_s = 1.5\text{equiv}$$

❖ Equivalent of a solute of a solute, Equiv_s is 1.5equiv

◦ Then;

$$\text{Normality, } N = \frac{\text{Equivalent of a solute, } \text{Equiv}_s}{\text{Volume of a solution, } V_{sn}}$$

$$N = \frac{1.5 \text{ Equiv}}{0.6L}$$

$$N = 2.5N$$

Try this

1. Calculate the equivalent weight of calcium hydroxide ($Ca(OH)_2$). Assume that all hydroxide ions are donated when it reacts with an acid.
2. Calculate the normality of a solution made by dissolving 60g of sodium hydroxide in water to form $500cm^3$ of a solution.

5.2 Standard Solutions

- Standard Solution is a solution of known concentration.
- The standard solution can be classified into two
 - a) ***Primary standard solution*** is the solution that is prepared by dissolving a known amount of primary standard in a specific volume of a known liquid. Example: Na_2CO_3 for titration.
- This solution may be titrated against acid in order to standardize a particular acid.

Characteristics of Primary standards solution

- i. High degree of purity.
- ii. Should not decompose with time.
- iii. Should not hygroscopic or efflorescent (no water content).
- iv. Not volatile.
- v. Should be high soluble.
- vi. Should have high molecular mass.

b) Secondary standard solution is the solution that is prepared from by standardization using primary standards solution.

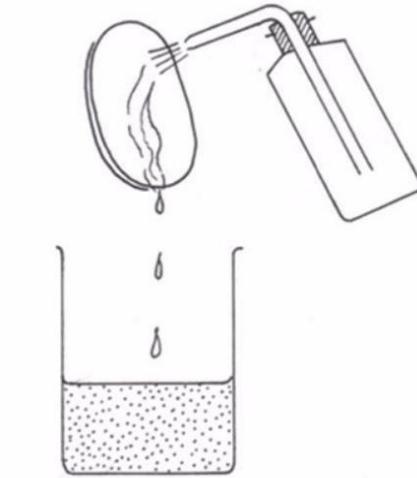
- Example;
- Dilute Hydrochloric acid (HCl)
- Dilute sulphuric acid (H_2SO_4)
- Dilute Sodium hydroxide (NaOH)

5.2.1 Preparation of base standard solution

- i. Carefully weigh the required mass of the substance
 - a) Weigh empty watch glass and note reading, M_1
 - b) Add the mass of substance to watch glass and Weigh watch glass together with mass of substance, M_2
 - c) Subtract mass of watch glass and substance from empty watch glass to get mass of a solute substance.

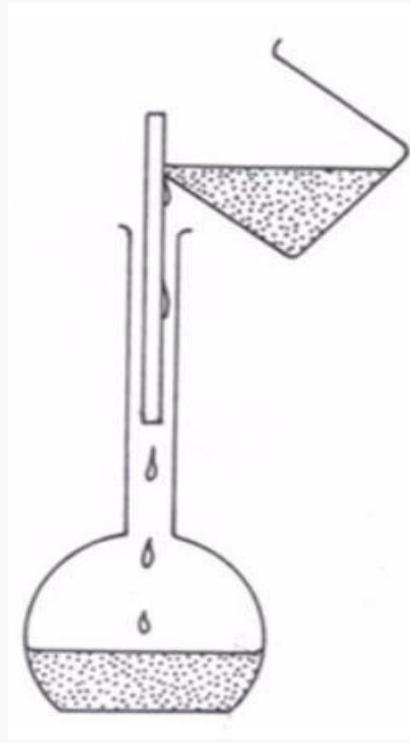
$$M_s = M_2 - M_1$$

- ii. Transfer the primary standard into a beaker. Rinse the watch glass using distilled water from wash bottle and pour rinse water into the beaker.



- iii. Stir the mixture in the beaker with glass rod until all the soluble is dissolved, then transfer the solution into a volumetric flask.
- iv. Rinse the beaker and the glass rod using distilled water from a wash bottle and pour rinse water into the volumetric flask.

v. Add water to the solution to just below the graduated mark on the volumetric flask.



vi. Top up the solution with water up to the graduated mark using a clean dropper to make sure that the lower meniscus is exactly at the graduated mark.

vii. Stopper the flask and invert it several times to ensure the solution is homogenous.

viii. Work out the molar concentration (Molarity) of the solution.

ix. Label the solution correctly by indicating its name and concentration.

Example: Sodium carbonate(Na_2CO_3)

Molarity, M = 0.1M

NOTE: This method can be used to standardize any chemical substance which is in solid form not only base(eg oxalic acid)

5.2.2 Preparation of acid standard solution

- Standard solution of mineral acids are prepared by dilution of the concentrated acids.
- Always ***acid is added into water*** and ***not*** vice versa.
 - i. Calculate molar concentration (Molarity) stock solution by using data indicated on reagent bottle.
- Data indicated on reagent bottle are:-
 - Density (specific gravity)
 - Assay (percentage purity)
 - Molecular mass, M_m /Molecular weight, M_w

$$M = \frac{P_p \times \rho \times 1000}{M_m}$$

Where: M = Molar concentration(molarity)

P_p = Percentage purity (assay)

ρ = Density(specific gravity)

M_m = Molar mass

- ii. Then, measure the volume of concentrated acid by using a measuring cylinder. The volume can be determined using the dilution law as explained in the next slide.

Dilution law

- In order to estimate the volume of the concentrated acid to be measured for preparation of standard solution one must use the dilution law.
- **Dilution law states that**, “The product of molarity and volume of concentrated solution is equal to the product molarity and volume of diluted solution.”

$$M_c V_c = M_d V_d$$

∴ This means

$$V_d = V_c + V_w$$

Where: M_c = molarity of concentrated solution

V_c = Volume of concentrated solution

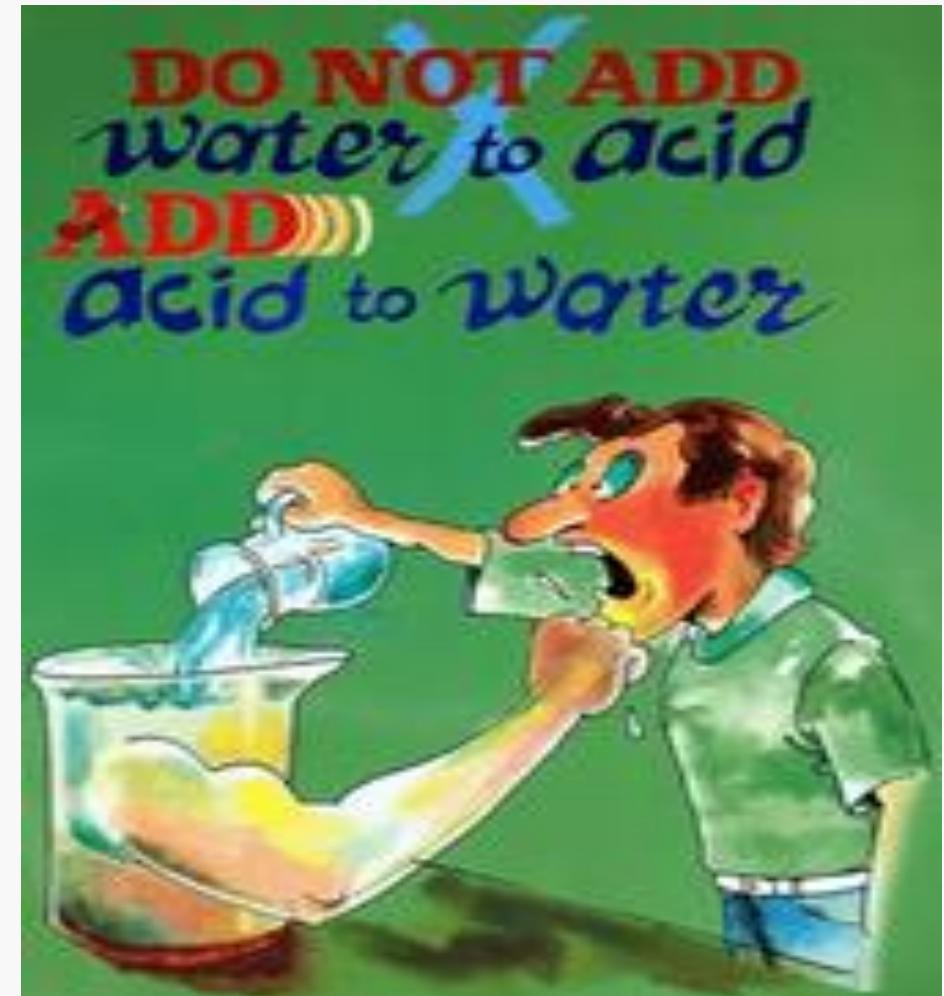
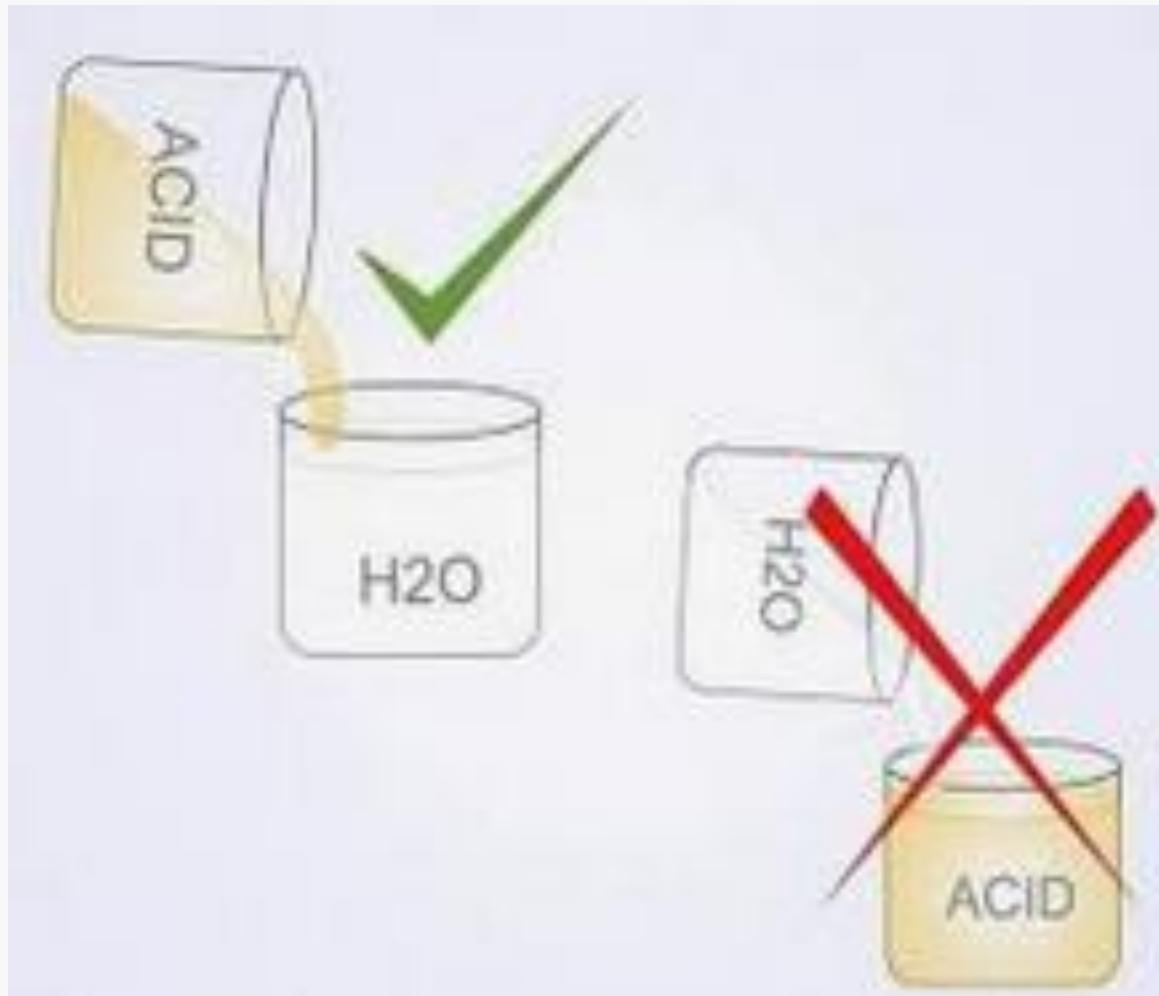
M_d = molarity of dilute solution

V_d = Volume of dilute solution

- V_w = Volume of water
- Dilution factor is the ratio of volume of concentrated volume of the ignition solution to the volume of diluted solution. It is the notation that shows how much of the original stock solution is present in the total solution, after dilution.

iii. Put water in the 250ml beaker about half of it and pour acid into water in the beaker.

Remember: Always add acid slowly into water and not vice versa.



- iv. Rinse the measuring cylinder with distilled water and add in the beaker.
- vi. Transfer the solution in the beaker into the volumetric flask. The beaker should be rinsed two times with water and pour water in the flask.
- vii. Add water to the solution to just below the graduated mark on the volumetric flask. Then fill up the flask with water to the lower meniscus of the graduation mark with clean dropper.

NOTE: This method can be used to standardize any chemical substance which is in liquid form not only acid(eg ammonia solution).

Example 1

- The industrial sulphuric acid is usually labeled as containing 1.84g/ml and percentage purity of 98%. Calculate the molarity of the acid.

Data given

Required: Molar concentration(molarity), M

Percentage purity (assay), $P_p = 98\% = \frac{98}{100}$

Density(specific gravity), $\rho = 1.84g/ml$

Molar mass, $M_m = (1x2) + 32 + (16x4) = 98g/mol$

Solution

$$M = \frac{P_p \times \rho \times 1000}{M_m}$$

$$M = \frac{98 \times 1.84 \times 1000}{100 \times 98}$$

$$M = 18.4M$$

- ❖ Molarity of the acid is 18.4M

Example 2

- A solution was prepared by diluting 5ml of 2M nitric acid to make one litre of a solution for titration experiment. Calculate
 - i) The molarity of the solution prepared.
 - ii) The volume of water that was added to prepare the solution in (a) above.

Data given

Molarity of concentrated solution, $M_c = 2M$

Volume of concentrated solution, $V_c = 5ml$

Volume of dilute solution, $V_d = 1L = 1000ml$

i) Required:

Molarity of dilute solution, M_d

$$2M \times 5ml = M_d \times 1000ml$$

$$M_d = \frac{2M \times 5ml}{1000ml}$$

$$M_d = 0.01M$$

❖ Molarity of dilute solution, M_d

is $0.01M$

ii) Required: Volume of water, V_w

$$V_d = V_c + V_w$$

$$V_w = V_d - V_c$$

$$V_w = 1000ml - 5ml$$

$$V_w = 995ml$$

❖ Volume of water, V_w is $995ml$

Try this

1. The industrial hydrochloric acid is labeled as containing $1.14\text{g}/\text{cm}^3$ and percentage purity of 35%.
 - a) Calculate the molarity of the acid.
 - b) What will be the volume of the solution made when 15cm^3 of this acid is diluted to form a solution of molarity 0.05M.
2. The percentage purity of nitric acid is 68% and it's density is $1.42\text{g}/\text{cm}^3$. 10cm^3 of this acid was diluted to 500cm^3 , calculate
 - a) Molarity of the acid in the bottle.
 - b) The molarity of the solution that was prepared.

5.2.3 Titration

- Titration is the process of adding one solution into another solution until the end point reaches.
- ***End point*** is the point during titration where the indicator shows a sudden colour change.
- During titration two solutions; one with known concentration and another with unknown concentration are used.
- A solution with unknown concentration which is undergoing analysis is known as ***analyte (titrate)***.
- A solution of known concentration which is used to analyse the solution of unknown concentration is called ***titrant***.

- **Equivalence point** (stoichiometric point) of a chemical reaction is the point at which chemically equivalent quantities of bases and acids have been mixed.
- This means, the moles of acid are equivalent to the moles of base.
- At equivalence point;

$$Equiv_a = Equiv_b$$

Since; $Equiv = NV$

Then: $N_a V_a = N_b V_b$

NOTE: *Equivalence point* and *end point* are quite different terms, but ideally equivalence point and end point of a chosen indicator *should be identical*.

Suitable indicator for titration

- Indicator is a substance which gives a visible colour change in presence or absence of a threshold concentration of a substance(acid or alkali).

OR

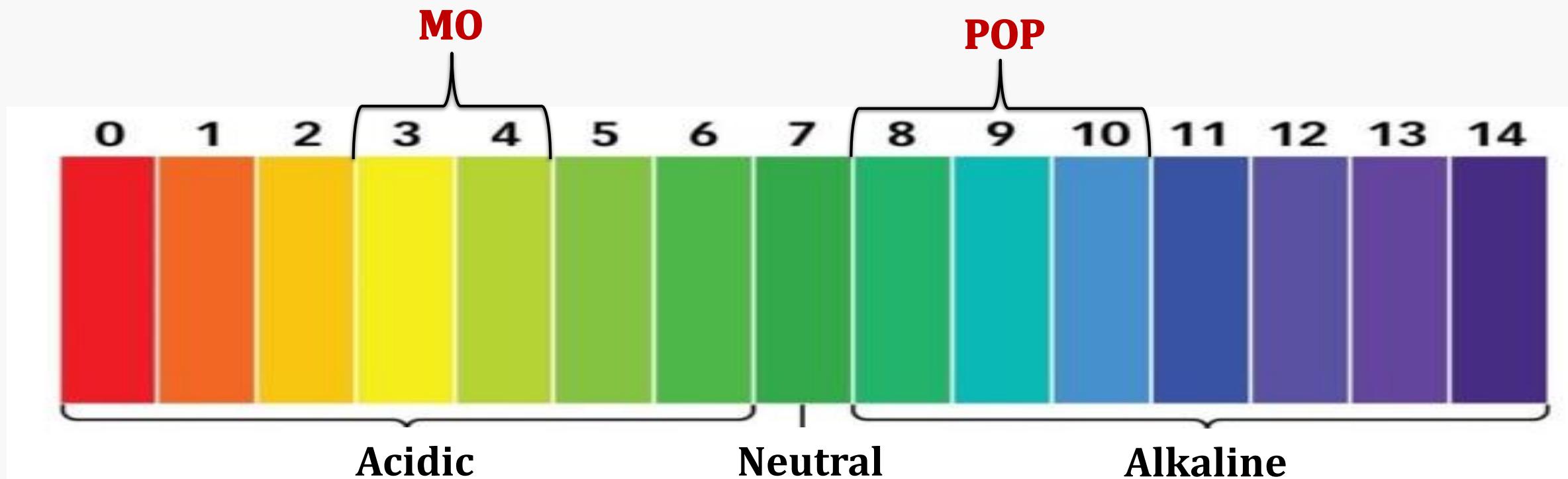
- Indicator is a substance that changes it's colour according to the chemical change(pH change) of a solution.
- There are three common indicators used in school laboratory are;
 - 1) Methyl orange(MO)
 - 2) Phenolphthalein (POP)
 - 3) Litmus

- The indicator to be used in titration depends on the type of an acid and a base used (strong or base).
- Each indicator has it's threshold pH in which change in colour occurs.
- The indicator which changes it's colour at high pH (more than 7) is suitable for the titration when ***base used is strong.***
- The indicator which changes it's colour at low pH (less than 7) is suitable for the titration when ***acid used is strong.***

NOTE: Litmus is not used in titration because the pH range over which it changes colour is too great.

End point pH of MO and POP

Indicator	pH range	Colour in Base	Colour in Acid
Methyl Orange(MO)	3 – 4.5	Yellow	Pink
Phenolphthalein(POP)	8 – 10	Pink	Colorless
Litmus	6 – 8	Blue	Red



- The following table shows the suitable indicators for a particular titration.

TITRATION		SUITABLE INDICATOR
BASE	ACID	
Strong	Strong	Any indicator
Weak	Strong	Methyl orange(MO)
Strong	Weak	Phenolphthalein(POP)
Weak	Weak	No indicator available

5.3 Volumetric Calculations

- In calculations involving volumetric analysis the following mathematical relationship is used;

$$\frac{M_a V_a}{n_a} = \frac{M_b V_b}{n_b}$$

This can also be written as;

$$\frac{M_a V_a}{M_b V_b} = \frac{n_a}{n_b}$$

Where;

M_a = Molarity of an acid

V_a = Volume of an acid

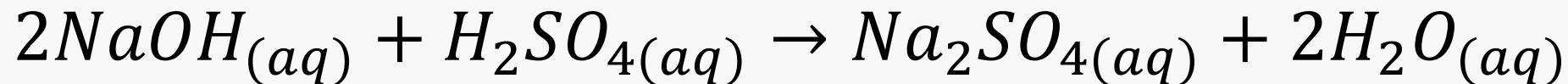
M_b = Molarity of a base

V_b = Volume of a base

n_a = Moles of an acid that react with a base

n_b = Moles of a base that react with an acid

NOTE: n_a and n_b can be obtained from the balanced chemical equation for the neutralization reaction.



$$n_b = 2 \quad n_a = 1$$

5.3.1 Standardization common mineral acids

- As most of mineral acids are secondary standards, they should be standardized by primary standards of base.

Example 1

1. 20ml of hydrochloric acid required 25ml of 0.1M sodium carbonate for complete neutralization.
 - a) Write the balanced chemical equation for the reaction.
 - b) Calculate the molarity of the acid.
 - c) Determine the concentration of the acid.

Data given:

Volume of acid, $V_a = 20ml$

Volume of base, $V_b = 25ml$

Molarity of base, $M_b = 0.1M$

Example 2

- 23cm^3 of 0.25molar sulphuric acid were able to neutralize 20cm^3 of potassium hydroxide. Calculate the concentration of KOH in;
 - a) Mole per litre.
 - b) Gram per litre.

Data given:

Volume of acid, $V_a = 23\text{cm}^3$

Volume of base, $V_b = 20\text{cm}^3$

Molarity of acid, $M_a = 0.25M$

Volumetric Practical Hints

Colours of MO indicator



In a base
(Yellow)

At the end point
(Orange)

In an acid
(Pink)

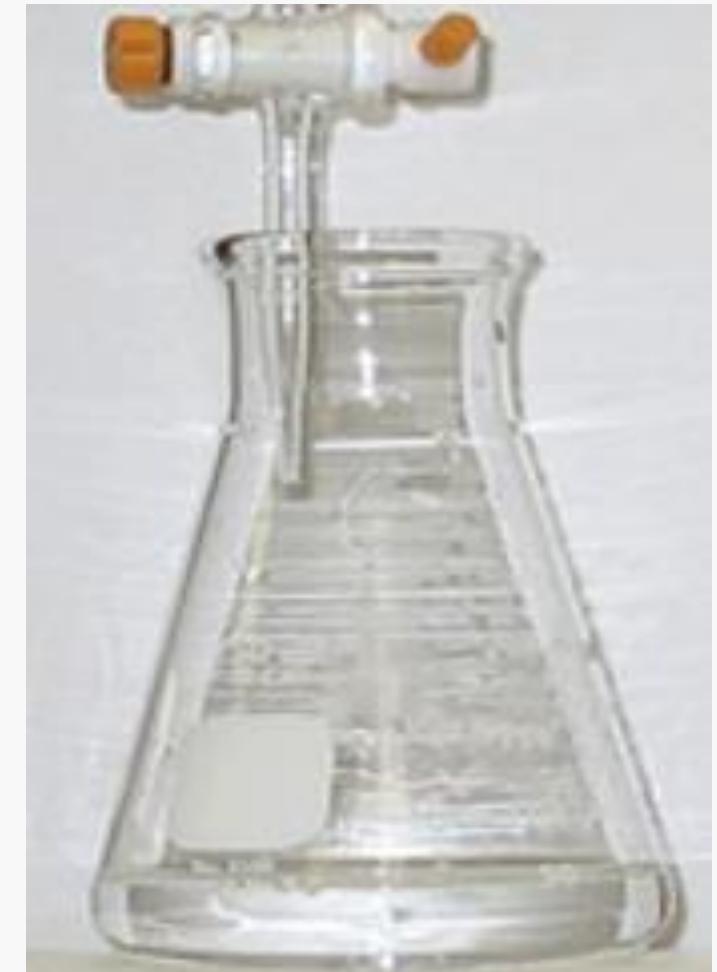
Colours of POP indicator



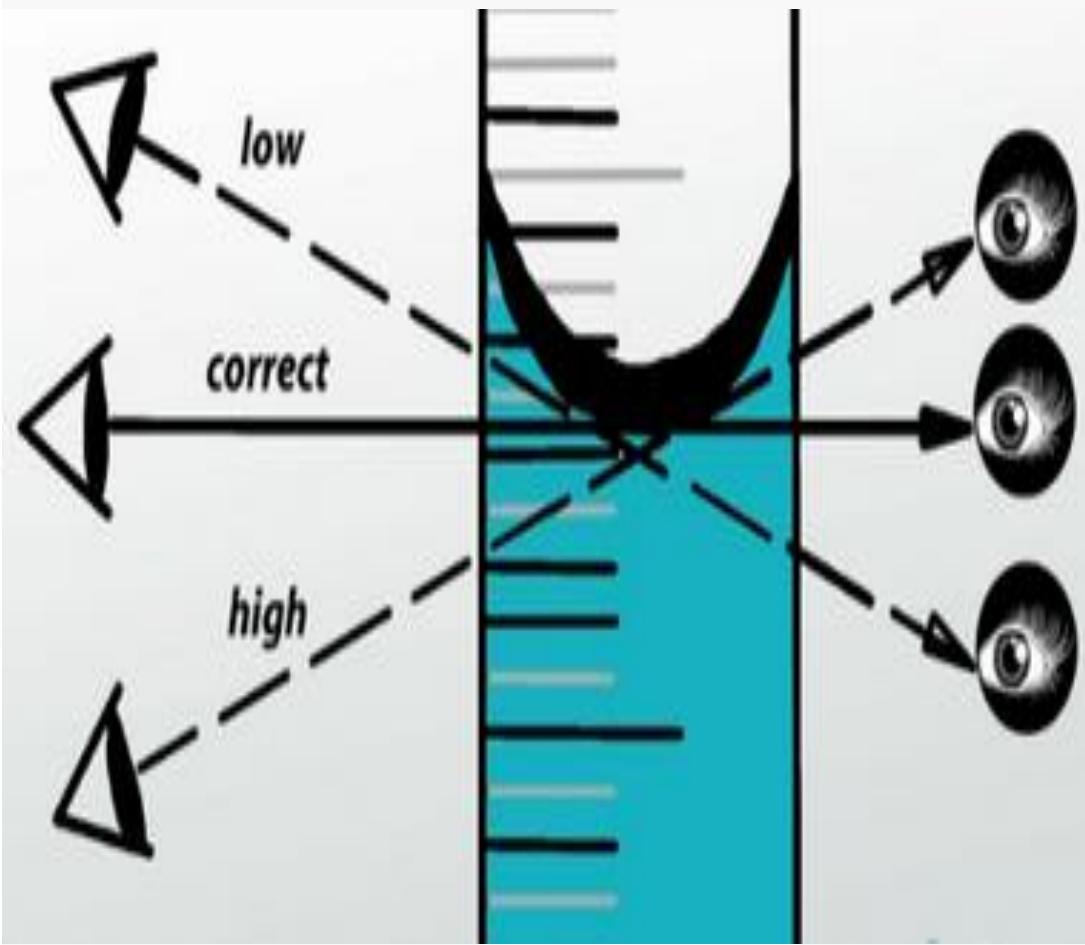
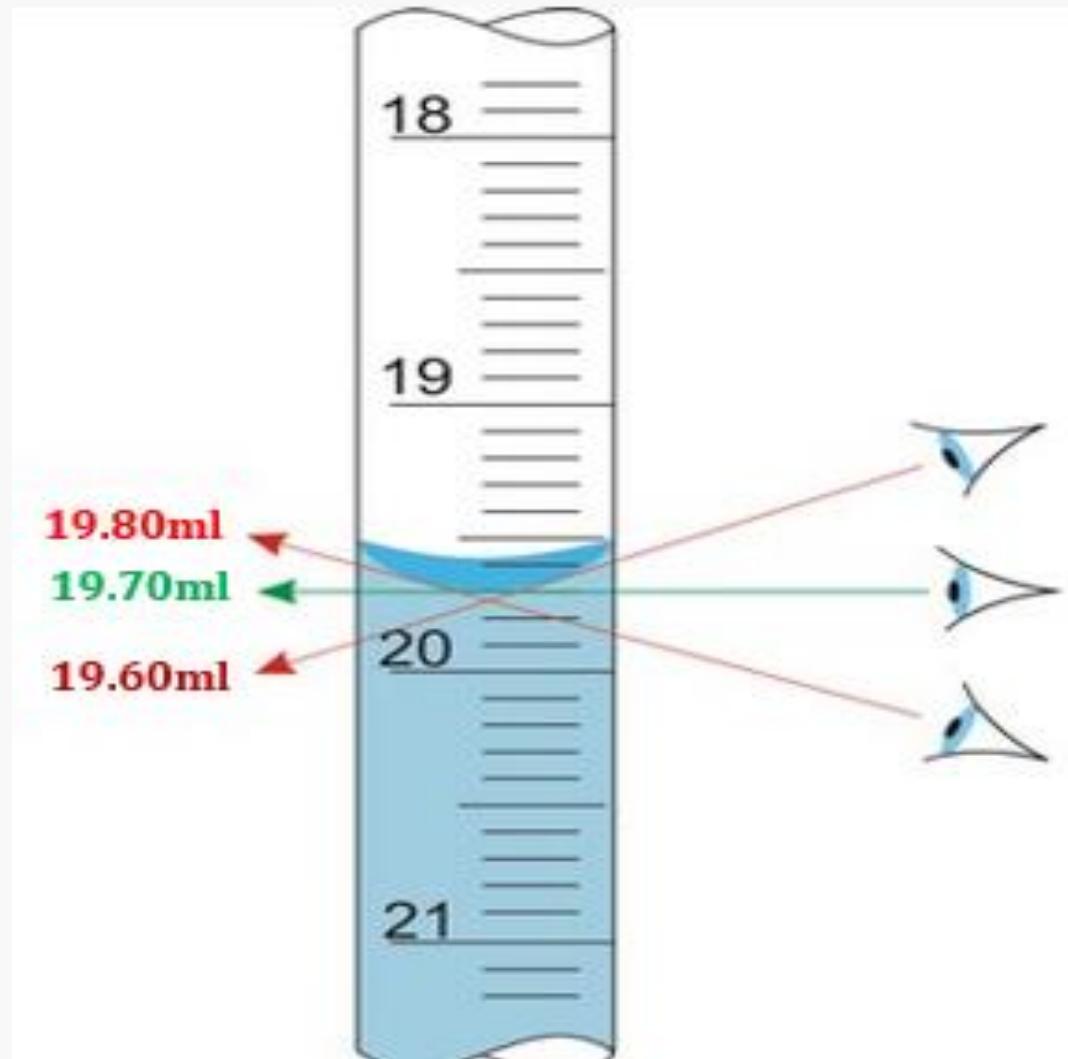
In a base
(Pink)



At the end point
(faint pink)



In an acid
(colourless)



Errors that can be made during titration

- i. ***Misjudging the colour*** of the indicator near the end point.
- ii. ***Misreading the volume*** due to parallax error or error in counting unmarked graduation marks.
- iii. ***Using contaminated solutions***
- iv. ***Using diluted solutions*** – This may happen if the burette and/or pipette was not rinsed with transferred solution after being rinsed with distilled water.
- v. ***Using dirty apparatus*** - if apparatus was not properly cleaned, it may be contaminated with previous reagents.

vi. Leaking burette - sometimes burettes leak slowly enough to allow titration, but will lose several tenths of millilitre if left for several minutes.

vii. Not transferring all the volume from the pipette - shaken pipette may lose a drop of the solution when it is being moved between flasks.

viii. Not filling burette properly - if there is an air lock in the burette stopcock it can block the flow of the titrant.

ix. Rinsing burette and/or pipette with wrong solution.

x. Losing solution and too vigorous swirling/shaking can end in liquid splashing from the titration flask before the end point had been reached.

Precautions to be taken during titration

- i. There should be no any leakage in the burette
- ii. The eye should be kept in level with liquid surface while taking pipette and burette readings.
- iii. Always read from lower meniscus at eye level.
- iv. An air bubble in the nozzle of the burette must be removed before taking the initial reading.
- v. Touch the inner surface of the titration flask to expel the last drop, but don't blow the pipette to do so.

- vi. Shaking of the titration flask should be continuous when adding solution from the burette. Don't shake vigorously you may lose the solution.
- vii. Do not waste your time to bring the burette to zero before each titration.
- viii. Addition of the indicator must be according to the procedure.
- ix. Use your index finger(finger next to thumb) while pipetting.

PRACTICAL ACTIVITY

5.3.2 Determination of the relative atomic mass of unknown element in an acid or alkali

- Here one must first find the molar mass (M_m) of a particular acid or alkali which is the same as relative molecular mass (RMM) in magnitude.
- The major difference is that M_m has units (g/mol) but RMM has no units.
- Since relative molecular mass of a compound is the sum of the relative atomic masses of all the elements in a compound, thus the atomic mass of an element can simply be obtained from it.

Example 1:

- The molar mass of a compound YOH is 56g/mol. Calculate the atomic mass of element Y and eventually name it.

Data given

- Molar mass, $M_m = 56\text{g/mol}$
- Atomic mass of oxygen, $A_O = 16$
- Atomic mass of hydrogen, $A_H = 1$
- Required: Atomic mass of Y, A_Y

Solution

$$RMM = A_Y + A_O + A_H$$

Since in magnitude RMM = M_m

$$56 = A_Y + 16 + 1$$

$$56 = A_Y + 17$$

$$A_Y = 56 - 17$$

$$A_Y = 39$$

- The atomic mass of element Y is **39**.
- The element Y is **potassium**.

Example 2:

- 2g of QOH were dissolved in water to form 250ml of a solution. The molarity of the resulting solution is 0.2M.

- a) Find the atomic mass of Q
- b) Name the element Q.

Data given

- Mass of a solute, $M_s = 2\text{g}$
- Volume of the solution, $V_{sn} = 250\text{ml} = 0.25\text{L}$
- Molarity, $M = 0.2\text{M}$
- Molar mass of QOH, M_m

Solution

$$M = \frac{M_s}{V_{sn}xM_m}$$
$$M_m = \frac{M_s}{V_{sn}xM} = \frac{2}{0.25x0.2}$$
$$M_m = 40$$

$$RMM = A_Q + A_O + A_H$$

Since in magnitude $RMM = M_m$

$$40 = A_Q + 16 + 1$$

$$A_Q = 40 - 17$$

$$A_Q = 23$$

- The atomic mass of element Q is **23**.
- The element Q is **sodium**.

Example 3:

- Solution Q was made by dissolving 2.65g of a M_2CO_3 in water and diluting it to $250cm^3$. $25cm^3$ of solution Q required $20cm^3$ of 0.25molar HCl solution.
- a) Calculate the concentration of solution Q in.
 - i. Mole per litre
 - ii. Gram per litre
 - b) Determine the relative atomic mass of **M** in M_2CO_3 and name the element **M**.

Data given

- Mass of a solute, $M_s = 2.65\text{g}$
- Volume of the solution, $V_{sn} = 250\text{ml} = 0.25\text{L}$
- Volume of base, $V_b = 25\text{cm}^3$
- Volume of acid, $V_a = 20\text{cm}^3$
- Molarity of acid, $M_a = 0.25\text{M}$
- Molar mass of M_2CO_3 , M_m

Solution

$$\frac{M_a V_a}{n_a} = \frac{M_b V_b}{n_b}$$



$$n_b = 2 \quad n_a = 2$$

$$M_b = \frac{M_a V_a n_b}{V_a n_a} = \frac{0.25 \times 20 \times 1}{25 \times 2}$$

$$M_b = 0.1\text{M}$$

The molarity of base $M_b = 0.1\text{mol/L}$

$$C_b = \frac{M_s}{V_{sn}} = \frac{2.65g}{0.25L}$$

$$C_b = 10.6g/L$$

From:

$$M_b = \frac{C_b}{M_m}$$

$$M_m = \frac{C_b}{M_b} = \frac{10.6g/L}{0.1mol/L}$$

$$M_m = 106g/mol$$

$$RMM = (A_M \times 2) + A_C + (A_O \times 3)$$

Since in magnitude $RMM = M_m$
 $\therefore 106 = (A_M \times 2) + 12 + (16 \times 3)$
 $106 = 2A_M + 12 + 48$
 $106 - 60 = 2A_M$
 $\frac{46}{2} = \frac{2A_M}{2}$
 $A_M = 23$

- The atomic mass of element M is **23**.
- The element M is **sodium**.

Try this

1. A solution was prepared by dissolving 5.3g of anhydrous sodium carbonate in water to make 2 litres of solution. It was observed that 25cm^3 of this solution reacted with 23cm^3 of a monoprotic acid of the formula HQ containing 1.971g of acid per litre of solution. Show all the details of your calculations; find the relative atomic mass of the element Q.
2. 3.552g of divalent metal hydroxide, A(OH)_2 was made up to 1000cm^3 . 25cm^3 of this solution required 30cm^3 of 0.04M sulphuric acid for complete neutralization. Find atomic mass of element A and name it.

4. A solution was prepared by dissolving 2.576g of a metal hydroxide YOH in water to make 500ml of a solution. 25cm³ of a prepared solution were against 0.1M hydrochloric acid solution. The volume of acid was recorded in the table below

EXPERIMENT	ROUGH	1	2	3
Final volume(cm ³)	23.90	28.10	22.90	46.00
Initial volume (cm ³)	0.00	05.00	0.00	23.00
Titre value (cm ³)				

- Write down the balanced chemical equation for the reaction.
- Calculate the molarity of the hydroxide solution.
- Calculate the relative molecular mass of YOH.

QUIZ

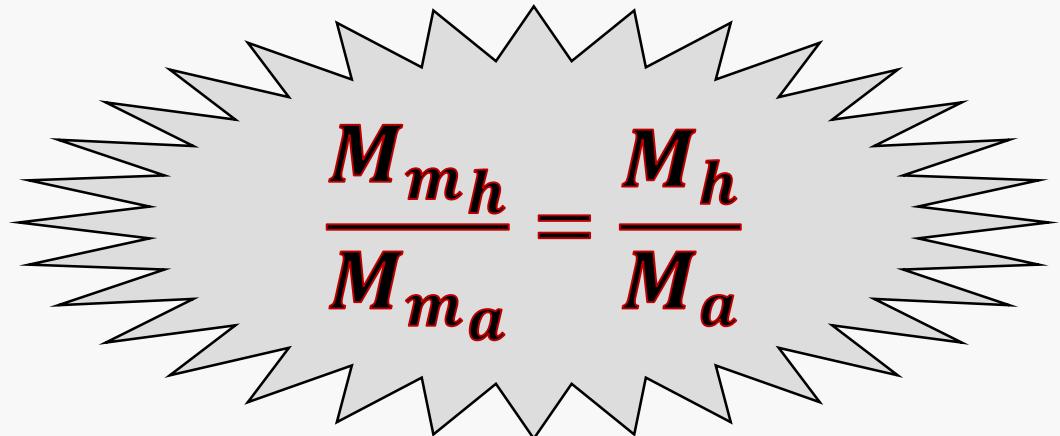
Time: 15Min

1. 2.8g of QOH were dissolved in water to form 250ml of a solution. The molarity of the resulting solution is 0.2M.
 - a) Find the atomic mass of Q
 - b) Name the element Q.
2. Solution Q was made by dissolving 2.65g of a M_2CO_3 in water and diluting it to 250cm³. 25cm³ of solution Q required 20cm³ of 0.25molar HCl solution.
 - a) Calculate the concentration of solution Q in.
 - i. Mole per litre
 - ii. Gram per litre
 - b) Determine the relative atomic mass of M in M_2CO_3 and name the element M.

5.3.3 Determination of the number of molecules of water of crystallization

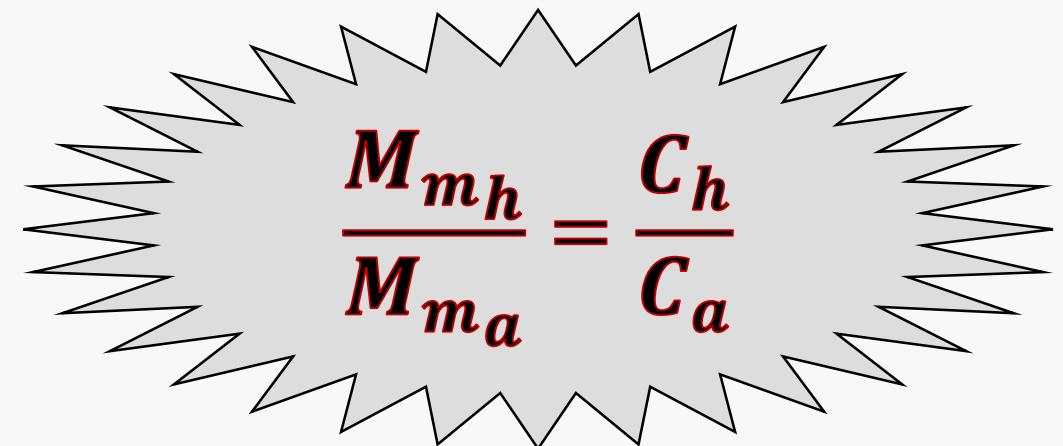
- **Crystal** is a solid that form by a regular repeated pattern of particles connected together. Crystallization:
- ***Crystallization is*** the process formation of solid crystals from a homologous solution
- The particles which are the building blocks of the crystalline solid can be atoms, molecules or ions.
- The intermolecular spaces of the crystals may be occupied by water which is known as **water of crystallization**.
- Water of crystallization is not covalently bonded to water molecules of the crystal. Thus, it can be removed when gently heated.

- The compound containing water of crystallization is called *hydrate*(hydrated compound)
- The compound that does not contain water in it is called *anhydrous*.
- The number of moles of water crystallization can be obtained as follows:



$$\frac{M_{mh}}{M_{ma}} = \frac{M_h}{M_a}$$

OR



$$\frac{M_{mh}}{M_{ma}} = \frac{C_h}{C_a}$$

Where:

Mm_h = Molar mass of hydrate

Mm_a = Molar mass of anhydrous

M_h = Mass of hydrate

M_a = Mass of anhydrous

C_h = Concentration of hydrate

C_a = Concentration of anhydrous

NOTE: One may find the molar mass of hydrate(M_{mh}) first as follows;

$$M_{mh} = \frac{M_h \times M_{ma}}{M_a}$$

OR

$$M_{mh} = \frac{C_h \times M_{ma}}{C_a}$$

Example 1

- Hydrated copper(ii) sulphate of mass 25g was gently heated and its mass decreased by 9g. Find the number of molecules of water of crystallization of the copper(ii) sulphate ($CuSO_4 \cdot XH_2O$).

Data given

- Mass of hydrate, $M_h = 25g$
- Mass of water, $M_w = 9g$

Solution

- $M_a = M_h - M_w$
- $M_a = 25 - 9 = 16g$
- From:

$$\frac{M_{mh}}{M_{ma}} = \frac{M_h}{M_a}$$

$$\frac{160 + 18x}{160} = \frac{25}{16}$$
$$x = 5$$

- The number of molecules of water of crystallization is 5.
- $\therefore CuSO_4 \cdot 5H_2O$

Example 2

- Exactly 15.8cm^3 of 0.3M HCl solution were enough to neutralize 23cm^3 of sodium carbonate solution made by dissolving 14.73g of hydrated sodium carbonate($\text{Na}_2\text{CO}_3 \cdot X\text{H}_2\text{O}$) in water to make 500cm^3 of a solution. Calculate
 - a) Molarity of base.
 - b) Number of molecules of water of crystallization in hydrated sodium carbonate.

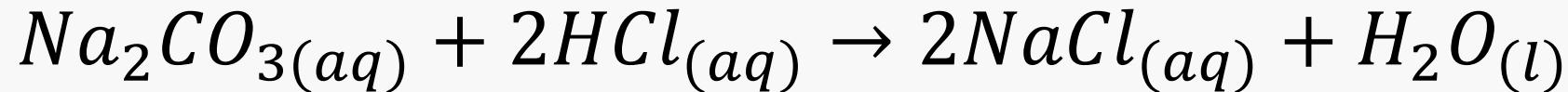
Data given:

Volume of base, $V_b = 15.8\text{cm}^3$ Molarity of acid, $M_a = 0.3\text{M}$

Mass of base solute, $M_{s,b} = 14.73\text{g}$ Volume of acid, $V_a = 23\text{cm}^3$

Volume of base solution, $V_{sn,b} = 500\text{cm}^3$

Solution



$$\frac{M_a V_a}{n_a} = \frac{M_b V_b}{n_b}$$

$$M_b = \frac{M_a V_a n_b}{V_b n_a} = \frac{0.3 \times 15.8 \times 1}{23 \times 2}$$

$$M_b = 0.103\text{M}$$

From:

$$M_{m_h} = \frac{C_h x M_{m_a}}{C_a}$$

▪ Concentration of hydrate, C_h

$$C_h = \frac{M_{s_b}}{V_{sn_b}} = \frac{14.73g}{0.5L}$$

$$C_h = 29.46\text{g/L}$$

▪ Concentration of anhydrous, C_a

$$C_a = M_b x M_m = 0.103x106$$

$$C_a = 10.918\text{g/L}$$

$$M_{m_h} = \frac{29.46x106}{10.918}$$

$$M_{m_h} = 286\text{g/mol}$$

Molar mass of $\text{Na}_2\text{CO}_3 \cdot X\text{H}_2\text{O}$ is 286.

$$RMM = A_{Na} x 2 + A_C + A_O x 3 + (A_H + A_O)x$$

Since in magnitude $RMM = M_m$

$$M_m = A_{Na} x 2 + A_C + A_O x 3 + (A_H x 2 + A_O)x$$

$$286 = 23x2 + 12 + 16x3 + (1x2 + 16)x$$

$$286 - 106 = 18x$$

$$\frac{180}{18} = \frac{18x}{18}$$

$$x = 10$$



Example 3

- A solution of 1.2g per 500ml of NaOH was treated against oxalic acid. 25ml of this solution required 15ml of oxalic acid for complete neutralization. The acid solution was made by dissolving 3.15g of hydrated oxalic acid ($H_2C_2O_4 \cdot XH_2O$) to make 500ml of a solution.
 - a) Write the balanced chemical equation for the reaction.
 - b) Find the molarity of base.
 - c) Find the molarity of acid.
 - d) Determine the value of X in $H_2C_2O_4 \cdot XH_2O$.

Data given

- Mass of a base solute, $M_{sb} = 1.71g$
- Volume of base solution, $V_{sn_b} = 500ml$
- Volume of base, $V_b = 25ml$
- Volume of acid, $V_a = 15ml$
- Mass of an acid solute, $M_{sa} = 3.15g$
- Volume of base solution, $V_{sn_a} = 500ml$

Solution

a) Req: Molarity of base, M_b

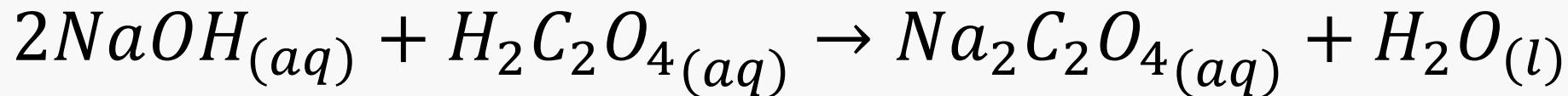
$$M_b = \frac{M_{sb}}{V_{sn_b} \times M_m b}$$

$$M_b = \frac{1.2g}{0.5L \times 40g/mol}$$

$$M_b = 0.06M$$

But from the formula

$$\frac{M_a V_a}{n_a} = \frac{M_b V_b}{n_b}$$



$$M_a = \frac{M_b V_b n_a}{V_a n_b}$$

$$M_a = \frac{0.06x25x1}{15x2}$$

$$M_a = 0.05M$$

❖ *Molarity of acid, M_a is $0.05M$*

From:

$$M_{m_h} = \frac{C_h x M_{m_a}}{C_a}$$

- Concentration of hydrate, C_h is given by:

$$C_h = \frac{M_{sa}}{V_{sn_a}} = \frac{3.15g}{0.5L}$$

$$C_h = 6.3\text{g/L}$$

- Concentration of anhydrous, C_a

$$C_a = M_a x M_{m_a} = 0.05 x 90$$

$$C_a = 4.5\text{g/L}$$

$$M_{m_h} = \frac{6.3 \times 90}{4.5}$$

$$M_{m_h} = 126 \text{ g/mol}$$

Molar mass of $H_2C_2O_4 \cdot 2H_2O$ is 126.

$$RMM = A_{Na} x 2 + A_C + A_O x 3 + (A_H + A_O)x$$

$$RMM = (2x A_H) + (2x A_C) + (4x A_O) + (2A_H + A_O)X$$

But since in magnitude RMM = M_m, then;

$$126 = (2x1) + (2x12) + (4x16) + (2x1 + 16)X$$

$$126 = 2 + 24 + 64 + 18X$$

$$126 - 90 = 18X$$

$$36 = 18X$$

$$\frac{36}{18} = \frac{18X}{18}$$

$$X = 2$$

❖ Thus, the formula hydrated acid will be $H_2C_2O_4 \cdot 2H_2O$

Alternatively:

- From the obtained molarity, M_a is $0.05M$ we can get the molar mass of hydrated acid, M_{m_h}

$$M_{m_h} = \frac{M_{s_a}}{V_{s_n} a \times M_a} = \frac{3.15}{0.5 \times 0.05} = 126$$

Molar mass, $M_{m_h} = 126g/mol$

$M_m(H_2C_2O_4 \cdot XH_2O) = 126g/mol$

$$RMM = (2xA_H) + (2xA_C) + (4xA_O) + (2A_H + A_O)X$$

But since in magnitude $RMM = M_m$, then;

$$126 = (2 \times 1) + (2 \times 12) + (4 \times 16) + (2 \times 1 + 16)X$$

$$126 = 2 + 24 + 64 + 18X$$

$$126 - 90 = 18X$$

$$36 = 18X$$

$$\frac{36}{18} = \frac{18X}{18}$$

$$X = 2$$

❖ Thus, the formula hydrated acid will be $H_2C_2O_4 \cdot 2H_2O$

Try this

1. The formula of hydrated ethandioic (oxalic) acid is $\text{H}_2\text{C}_2\text{O}_4 \cdot X\text{H}_2\text{O}$. The 12.6g of the acid was dissolved in 1000cm³ of solution. 12.5cm³ of this solution was found to neutralize 25.0cm³ of 0.1M sodium hydroxide solution. Calculate:

a) Concentration of the anhydrous acid in:

i. Mole/dm⁻³

ii. g/dm⁻³

b) Mass concentration of hydrated acid.

c) The value of x in the hydrated acid.

2. Hydrated sodium carbonate ($\text{Na}_2\text{CO}_3 \cdot \text{YH}_2\text{O}$) solution was made by dissolving 179g of its crystals to make a litre of solution. 20cm^3 of the solution was found to neutralize 25cm^3 of a molar solution of nitric acid. Show all procedures to find the value of Y in the formula of hydrated sodium carbonate.
3. A sample of mass 28.6g of hydrated sodium carbonate ($\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$) was heated such that it's water was entirely absorbed by 32g of anhydrous copper(ii) sulphate to blue compound ($\text{CuSO}_4 \cdot \text{XH}_2\text{O}$). Find the value of X in the copper(ii) sulphate.

4. 33.92g of hydrate sodium carbonate crystals were made up to 1000cm³ of solution. 25cm³ of this solution neutralized 30cm³ of 0.2M hydrochloric acid. Showing all the steps, calculate the number of moles of water of crystallization present in the hydrated crystals.

a) Concentration of the anhydrous base in:

i. Moledm⁻³

ii. gdm⁻³

b) Mass of water in each 33.92g of the hydrated base.

c) The value of x in the hydrated base.

5.3.4 Determination of the percentage purity of an acid or an alkali

- When a chemical substance (acid or base) is contaminated with unwanted materials it is said to be ***impure***.
- The unwanted materials which contaminate base or acid are called ***impurities***.
- If the impurities are removed, then an acid or base is said to be ***pure***.
- During chemical reaction *only pure sample* is said to be involved in the chemical reaction, but the *impurities remain unreacted*.

- The percentage purity and impurity of an acid or base can be determined as follows;

$$P_p = \frac{M_p}{M_{im}} \times 100\%$$

◻ Also:

$$P_i = \frac{M_i}{M_{im}} \times 100\%$$

- But, note that;

$$P_p + P_i = 100\%$$

❖ **Where**

P_p = Percentage purity

P_i = Percentage impurity

M_p = Mass of pure sample

M_i = Mass of impurities

M_{im} = Mass of impure sample

- Instead of mass one can use concentration.
- Hence the formula may be written as:

$$P_p = \frac{C_p}{C_{im}} \times 100\%$$

◻ **Also:**

$$P_i = \frac{C_i}{C_{im}} \times 100\%$$

❖ **Where**

C_p = Conc of pure sample

C_i = Conc of impurities

C_{im} = Conc of impure sample

Example 1:

Solution

- The magnesium sulphate weighing 15g is impure sample. If 3g of the sample are the contaminants, find the percentage purity of the sample.

Data given

Required: Percentage purity, P_p

Mass of impurities, $M_i = 3g$

Mass of impure sample, $M_{im} = 15g$

Mass of pure sample, M_p will be

$$M_p = M_{im} - M_i$$

$$M_p = 15g - 3g$$

$$M_p = 12g$$

$$\therefore P_p = \frac{M_p}{M_{im}} \times 100\%$$

$$P_p = \frac{12g}{15g} \times 100\%$$

$$P_p = 80\%$$

Percentage purity is 80%

Example 2

◦ A solution was prepared by dissolving 5.2g of impure NaOH in water to make 500ml of a solution. If 23ml of this solution required 25ml of 0.2M HCl. Calculate

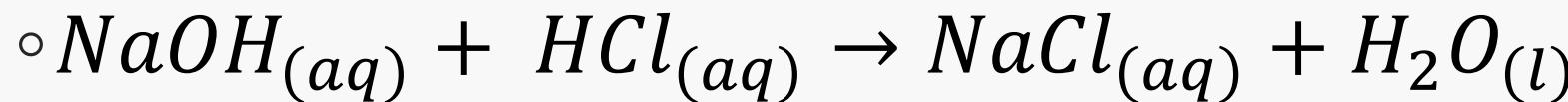
- Molarity of NaOH
- Percentage purity and impurity of NaOH.

Data given

- Mass of a base solute, $M_{sb} = 5.2\text{g}$
- Volume of base solution, $V_{snb} = 500\text{ml}$
- Molarity of base, $M_b = 0.2M$
- Volume of base, $V_b = 23\text{ml}$
- Volume of acid, $V_a = 25\text{ml}$

Solution

a) Required: molarity of base (NaOH)



$$\frac{M_a V_a}{n_a} = \frac{M_b V_b}{n_b}$$

$$M_b = \frac{M_a V_a n_b}{V_b n_a} = \frac{0.2 \times 25 \times 1}{23 \times 1}$$

$$M_b = 0.217\text{M}$$

◦ Molarity of base (NaOH), M_b is 0.217M

◦ b) Required: Percentage purity

◦ From: $P_p = \frac{C_p}{C_{im}} \times 100\%$

◦ Concentration of pure base, $C_p = M_m \times M_b = 40 \times 0.217$

$$C_p = 8.68 \text{ g/L}$$

◦ Concentration of impure base, $C_{im} = \frac{M_{sb}}{V_{sn_b}} = \frac{5.2 \text{ g}}{0.5 \text{ L}} = 10.4 \text{ g/L}$

$$C_{im} = 10.4 \text{ g/L}$$

◦ Then; $P_p = \frac{8.68}{10.4} \times 100\% = 83.5\%$

◦ Percentage purity is 83.5

◦ Percentage impurity = $100\% - 83.5\% = 16.5\%$

Try this

1. 25cm³ of impure sulphuric acid containing 5.2g/dm³ reacted with 25cm³ of sodium hydroxide solution made by dissolving 2.0g NaOH in distilled water to make 0.5 litre solution.
Calculate the percentage:
 - a) Purity of the acid.
 - b) Impurity of the acid
2. 8.50g of a sample of iron required just 75cm³ of 3.00M HCl to dissolve it and give a neutral solution. Calculate the percentage purity of the sample of iron.

3. Hydrochloric acid containing 25.5g of impure acid in a litre of a solution was titrated against 25cm³ of 0.40M sodium hydroxide solution. The volume of acid obtained was recorded in tabular form as follows:

EXPERIMENT	PILOT	1	2	3
Final volume(cm ³)	25.60	30.10	40.20	25.00
Initial volume (cm ³)	0.00	05.00	15.00	0.00
Volume used (cm ³)				

Calculate:

- Average volume of acid used
- Showing all your calculations clearly, calculate the percentage purity of an acid.

5.4 Application of Volumetric Analysis

5.4.1 Application of volumetric analysis in real life situations

- i. ***Medicine:*** In precipitation of solution and suspension in the hospitals for injection, volumetric analysis is mostly applied.
- ii. ***Waste disposal:*** The wastes such as industrial wastes may contain acid or base in them. They should be neutralized to make them harmless before they are disposed to the environment.
- iii. ***Agriculture:*** Determination of amendments that can be used for correcting soil pH applies volumetric analysis.

iv. *In industries:* Volumetric analysis is applied in various ways in industries such as:

- Determination of amount of alcohol in beers
- Determination of sucrose concentration in sugar cane

v. *Water analysis:* Used to determine hardness of water and determination of certain harmful elements or compound.

vi. *Standardization :* Volumetric analysis (titration) knowledge helps in the preparations of standard solutions of acids and bases.

5.4.2 Compare industrial and laboratory skills of volumetric analysis

TOPIC 6

IONIC THEORY AND ELECTROLYSIS

6.1 IONIC THEORY

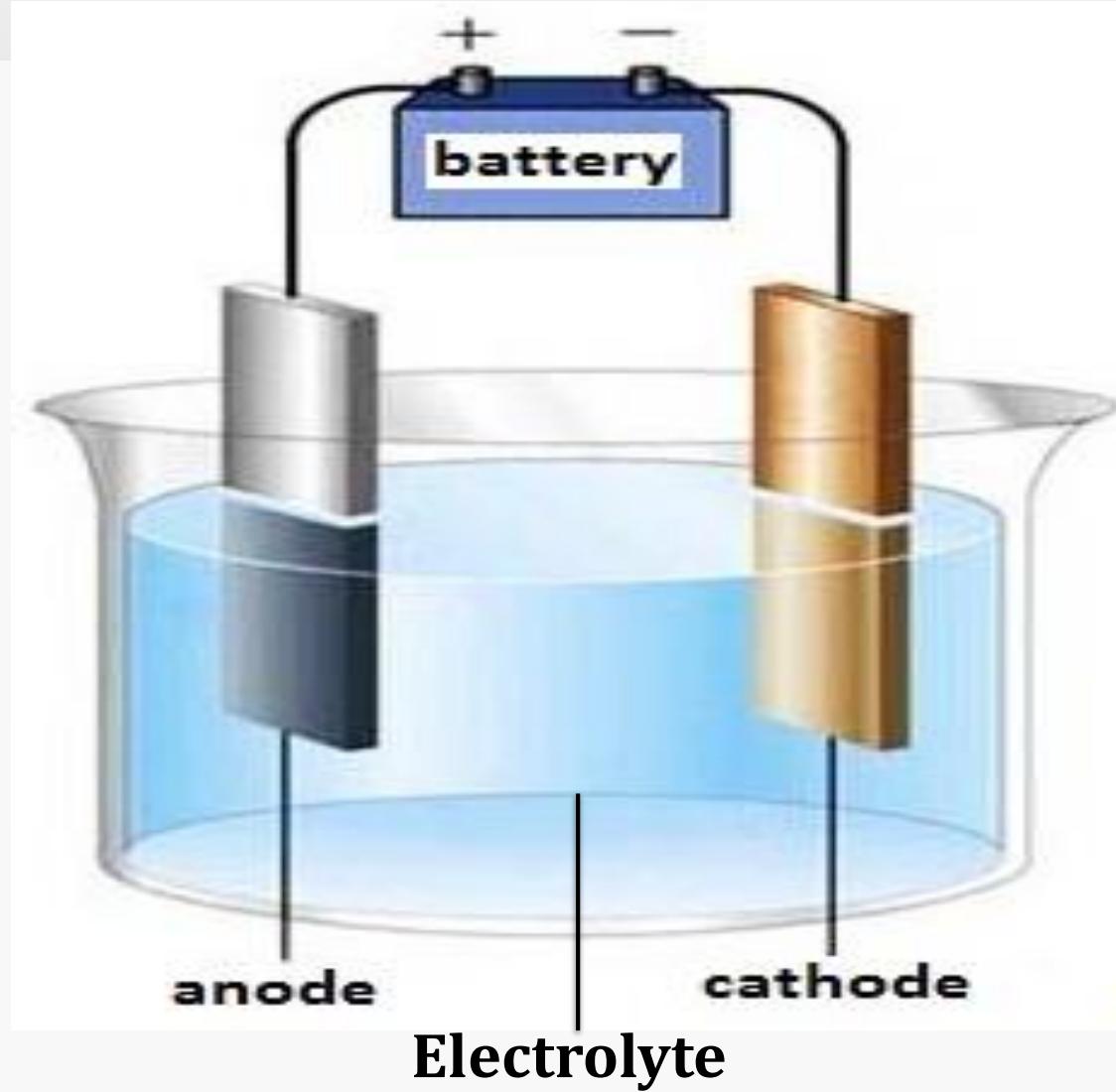
6.1.1 Introduction

- Electrolysis is the process in which an electric current is passed through the electrolytes causing the Chemical Reaction.

OR

- Electrolysis is the decomposition of an electrolyte by passing an electric current through it.
- The electrolytic cell called **voltammeter** is used for electrolysis.

The voltammeter



Terms used in electrolysis

- i) **Electrodes** are the poles or plates which enable electric current to enter and leave the electrolytes. They are made up of carbon or metal.
- ii) **Cathode** is the electrode which the electric current leaves the electrolyte. It is connected to the negative plate of the battery.
- iii) **Anode** is the electrode which the electric current enters the electrolyte. It is connected to the positive plate of the battery.
- iv) **Inert Electrode** is electrode which do not undergo any changes during electrolysis. For example platinum and carbon (graphite).

v) Ion is a charged atom or group of atoms.

vi) Cation is a positively charged ion.

vii) Anion is a negatively charged ion.

6.1.2 Electrolytes and non electrolytes

- **Electrolyte** is the solution or molten compound which allow electric current to pass through it. In solution/molten the ions are free to move, hence conduct electricity.
- It may be solution of Acids, Bases or Salts.
- **Non-electrolytes** is the solution or molten compound which electric current cannot pass through it. Eg: Ethanol, ester, benzene, sucrose(sugar).
- **Conductor** is a substance which allows electricity to pass through.
Example: Metals – Conducts by means of free electrons.
 Electrolytes – Conducts by means of ions and electrons.
- **Non-conductor (insulator)** is a substance which does not allow electricity to pass through. Example: Woods, plastic materials.

6.1.3 Types of Electrolytes

- There are two types of electrolytes includes
 - i. Weak Electrolyte
 - ii. Strong Electrolyte

i) Weak Electrolyte

- Weak electrolytes are the one which dissociates partially to give free ions in aqueous solution.
- Very few free ions are produced by weak electrolytes.
- Example:
- Phosphoric acid: $H_3PO_{4(l)} \rightleftharpoons 3H_{(aq)}^+ + PO_{4(aq)}^{3-}$
- Carbonic acid: $H_2CO_{3(l)} \rightleftharpoons 2H_{(aq)}^+ + CO_{3(aq)}^{2-}$
- Ethanoic acid: $CH_3COOH_{(l)} \rightleftharpoons CH_3COO_{(aq)}^- + H_{(aq)}^+$
- Ammonium hydroxide: $NH_4OH_{(l)} \rightleftharpoons NH_{4(aq)}^+ + OH_{(aq)}^-$
- Water: $H_2O_{(l)} \rightleftharpoons H_{(aq)}^+ + OH_{(aq)}^-$

ii) Strong Electrolyte

- Strong electrolytes are the one which dissociate completely into free ions in aqueous solution.

Example

- Sodium Chloride: $NaCl_{(s)} \rightarrow Na_{(aq)}^+ + Cl_{(aq)}^-$
- Sulphuric Acid: $H_2SO_{4(l)} \rightarrow 2H_{(aq)}^+ + SO_{4(aq)}^{2-}$
- Hydrochloric Acid: $HCl_{(l)} \rightarrow H_{(aq)}^+ + Cl_{(aq)}^-$
- Nitric Acid: $HNO_{3(l)} \rightarrow H_{(aq)}^+ + NO_{3(aq)}^-$
- Potassium hydroxide: $KOH_{(s)} \rightarrow K_{(aq)}^+ + OH_{(aq)}^-$

The ionic theory states that:

- *Ionic compound is composed with ions which are positively (cation) and negatively charged (anion) atoms or radicals.*
- These ions carry electric current through an electrolyte.
- **In solid compound**: Ions are not mobile due to strong force of attraction between Cation and anion
- **In molten compound**: Ions are free due to weaker force of attraction between cation and anion , hence conduct electricity.
- **In solution compound**: The compound dissociates into ions and the ions are free to move within the solution, hence conduct electricity.

Try this:

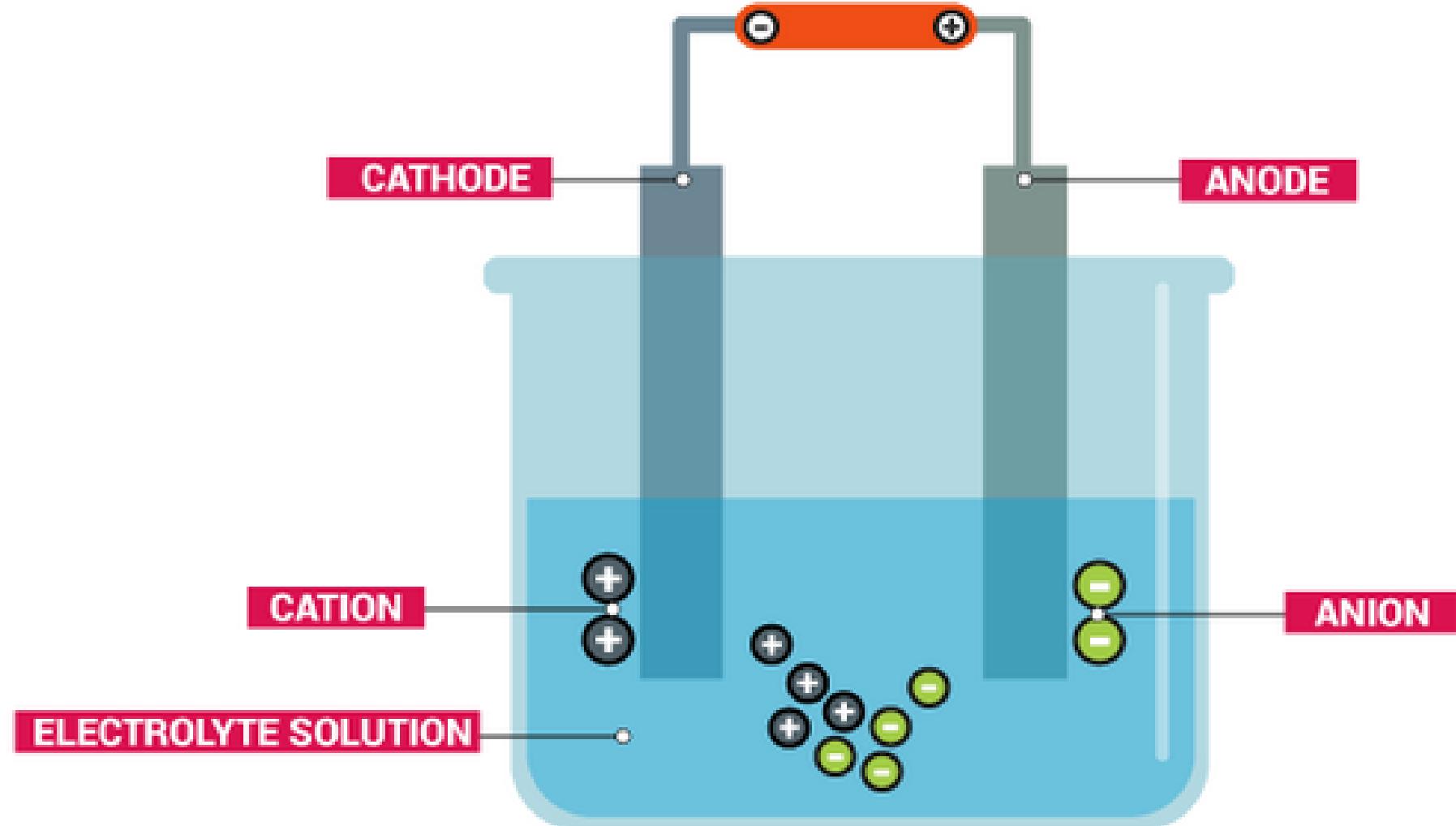
- i) Explain why Molten copper is not electrolyte but conduct electricity.
- ii) Explain why solution of sugar cannot conduct electricity.
- iii) Solid sodium chloride does not conduct electricity but aqueous sodium chloride does. Explain why?

6.2 MECHANISM OF ELECTROLYSIS

6.2.1 Migration of ions

- When an electric current is passed through an electrolyte in an electrolytic cell the free ions move to the electrodes,
- The movement of an ions to the electrolyte is referred to as migration,
 - Cation migrates towards the cathode (cations – cathode ions)
 - Anions migrates towards anodes (anions – anode ions)

Migration of ions during electrolysis



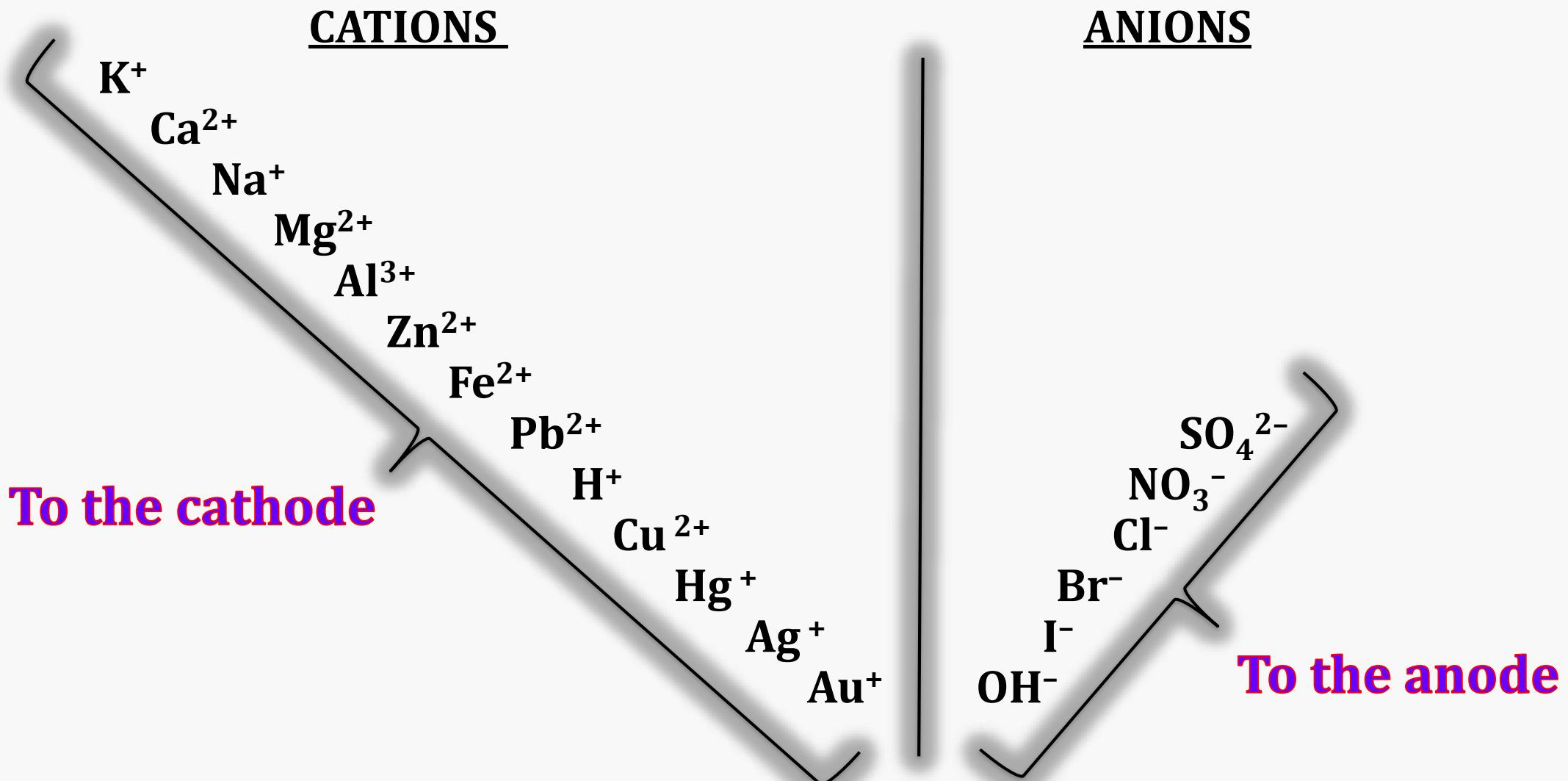
6.2.2 Factors Affecting selective Discharge of Ions

- At the electrodes, some of the ions gain or lose electrons; this has the effect of discharging the ions.
- 1) Position an ion in the electrochemical series
 - 2) The concentration of the ions
 - 3) The nature of the electrode

i) Position an ion in the electrochemical series

- Electrochemical Series is an arrangement of ions depending on their ease in discharging.
- More than one ion may be at the same electrode in the condition, but only one of them is *preferentially selected for discharge.*
- The ion that is found at the bottom of the series tends to discharge on preference to those found at the top.

The electrochemical series



Example

- Electrolysis of copper (II) sulphate solution with graphite electrodes.

◻ Dissociation of the ions:

- $\text{Cu SO}_4\text{(s)} \rightarrow \text{Cu}^{2+}\text{(aq)} + \text{SO}_4^{2-}\text{(aq)}$
- $\text{H}_2\text{O(l)} \rightleftharpoons \text{H}^+\text{(aq)} + \text{OH}^-\text{(aq)}$

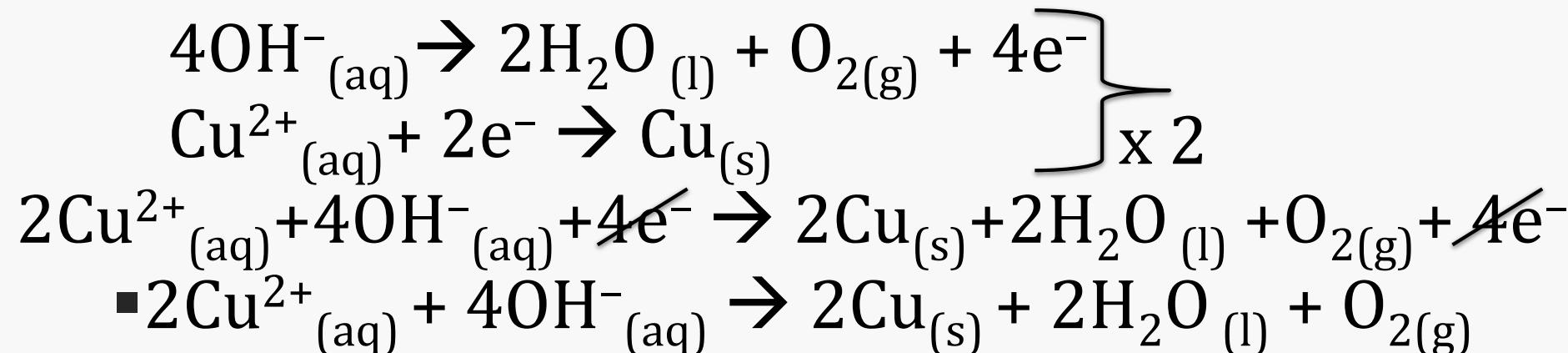
◻ At anode

- Ions present are SO_4^{2-} and OH^-
- But OH^- is found at the lower position than SO_4^{2-} in the electrochemical series so it is preferentially selected for discharge.
- $4\text{OH}^-\text{(aq)} \rightarrow 2\text{H}_2\text{O(l)} + \text{O}_{2(g)} + 4\text{e}^-$
- ❖ Oxygen gas is produced at anode

□ At cathode:

- Ions present are Cu^{2+} and H^+
- But Cu^{2+} is found at the lower position than H^+ in the electrochemical series so it is preferentially selected for discharge.
- $\text{Cu}^{2+}_{(\text{aq})} + 2\text{e}^- \rightarrow \text{Cu}_{(\text{s})}$
 - ❖ Solid copper is deposited at the cathode

□ Overall discharge equation: Combine the equations at anode & cathode



Conclusion:

- Discharge of OH^- disturbs the ionic equilibrium of water, hence more water ionizes to compensate the loss of OH^- ions in the solution. Excess H^+ produced increases acidity.
- The resulting solution will be more acidic due to the presence of H^+ ions and SO_4^{2-} ions remains in the solution forming sulphuric acid.
- The blue colour of the solution will *disappear* due to the discharge of copper.

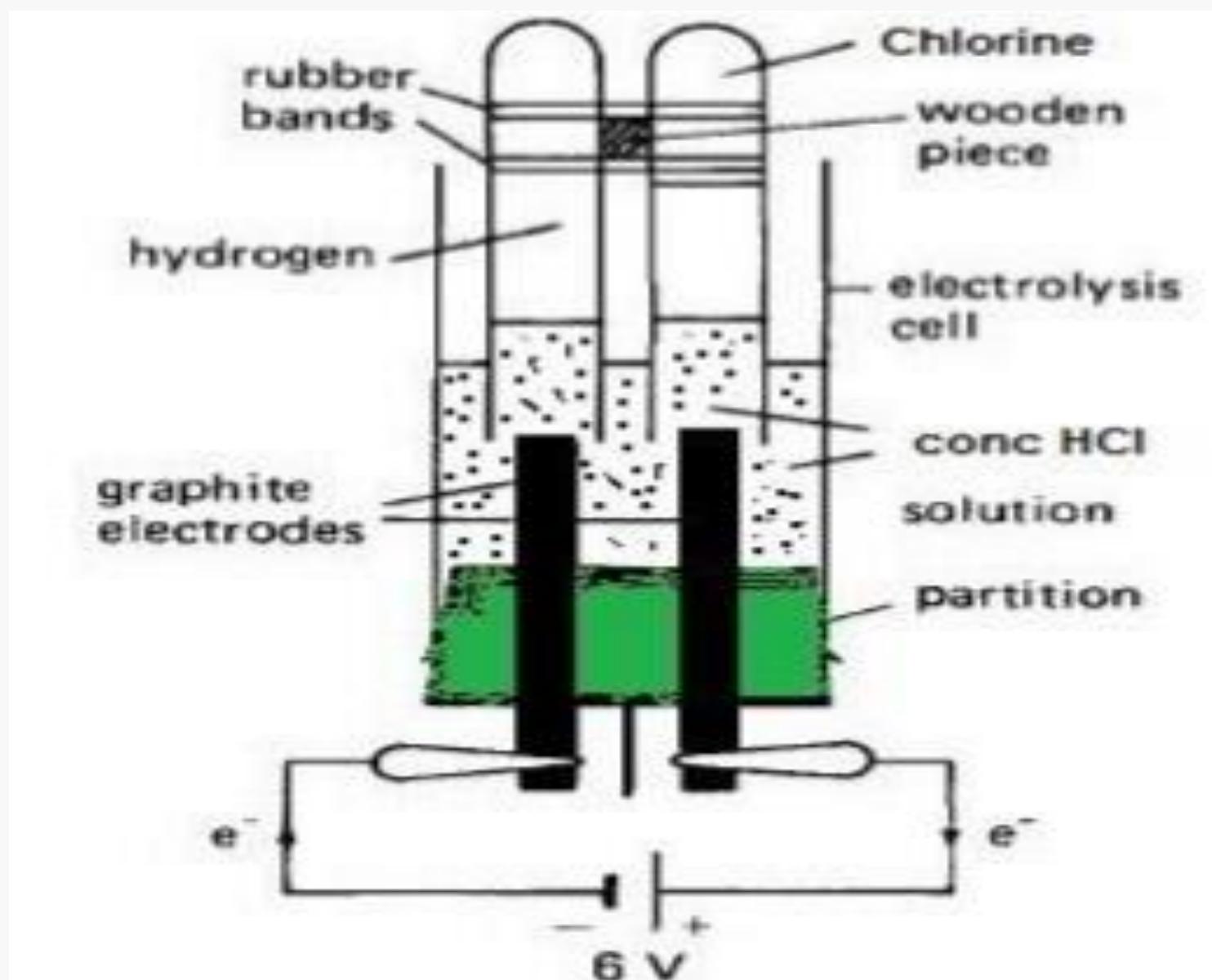
ii) Concentration

- A high concentration of ions tends to favour their discharge.
- This fact is most vivid when halides are considered.
- When a concentrated solution containing halide is electrolysed, halide will discharge on preference to OH^- .

Example

- Electrolysis of the concentrated solution of Hydrochloric Acid with Carbon Electrode

Diagram



Dissociation

- $\text{HCl}_{(\text{aq})} \rightarrow \text{H}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$
- $\text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{H}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$

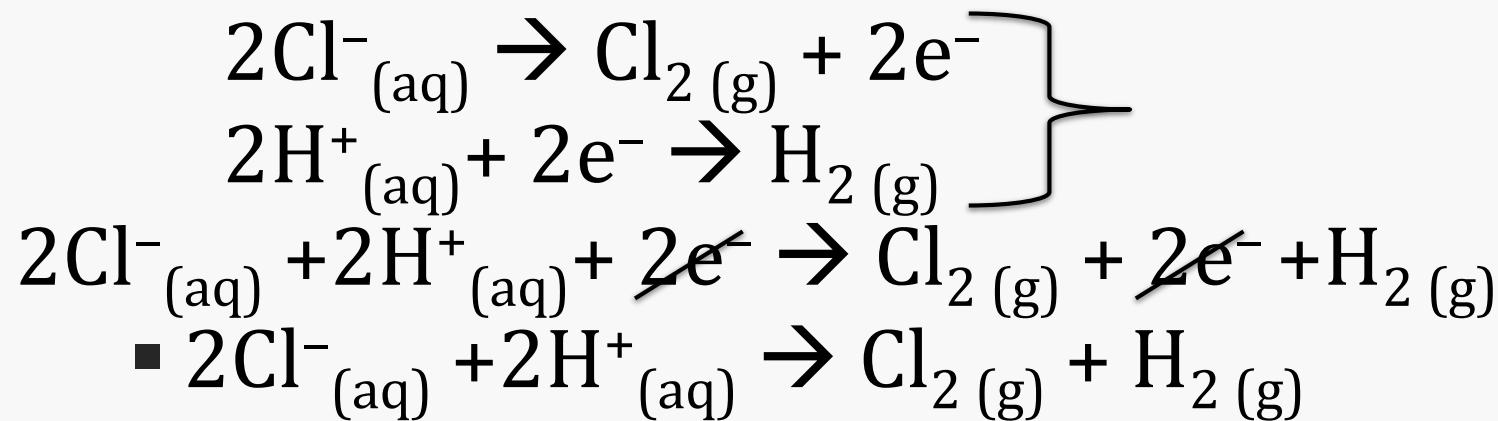
At anode

- Ions present are Cl^- and OH^-
 - But Cl^- is in higher concentration than OH^- in the electrolyte, so it is preferentially selected for discharge.
 - $2\text{Cl}^-_{(\text{aq})} \rightarrow \text{Cl}_2_{(\text{g})} + 2\text{e}^-$
 - Chlorine gas is produced at anode
- ❖ **NOTE:** However, if acid is very dilute some discharge of OH^- will also occur (ie mixture of Cl_2 & O_2 will be produced). As acid is diluted, there will not be a point at which production of Cl_2 ceases and replaced by O_2 .

At cathode:

- Ions present are H⁺ only.
- Thus, H⁺ will discharge.
- $2\text{H}^+_{(\text{aq})} + 2\text{e}^- \rightarrow \text{H}_2_{(\text{g})}$
 - ❖ Hydrogen gas is produced at the cathode

Overall discharge equation: Combine the equations at anode & cathode



iii) Nature of electrode used

- The Nature of electrode used, also Determines which ions are discharged.
- Here we have two important cases
 - a) Metal used as electrode for the electrolysis of the electrolyte containing its ions.
 - b) Mercury used as electrode for the electrolysis of an electrolyte containing sodium ions.

i) Metal used as electrode for the electrolysis of the electrolyte containing its ions.

- When the metal is used as electrode for electrolysis of the electrolyte containing its ions, the anode will dissolve and form ions in the electrolyte.

Example

- Electrolysis of copper (II) sulphate solution using copper electrodes.

◻ Dissociation of the ions:

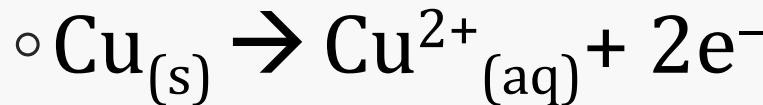
- $\text{Cu SO}_4\text{(s)} \rightarrow \text{Cu}^{2+}\text{(aq)} + \text{SO}_4^{2-}\text{(aq)}$
- $\text{H}_2\text{O(l)} \rightleftharpoons \text{H}^+\text{(aq)} + \text{OH}^-\text{(aq)}$

◻ At anode

- Ions present are SO_4^{2-} and OH^- . With this anode, there are three possibilities

1. Conversion of $\text{Cu}_{(\text{s})}$ to $\text{Cu}^{2+}_{(\text{aq})}$
2. Discharge of OH^-
3. Discharge of SO_4^{2-}

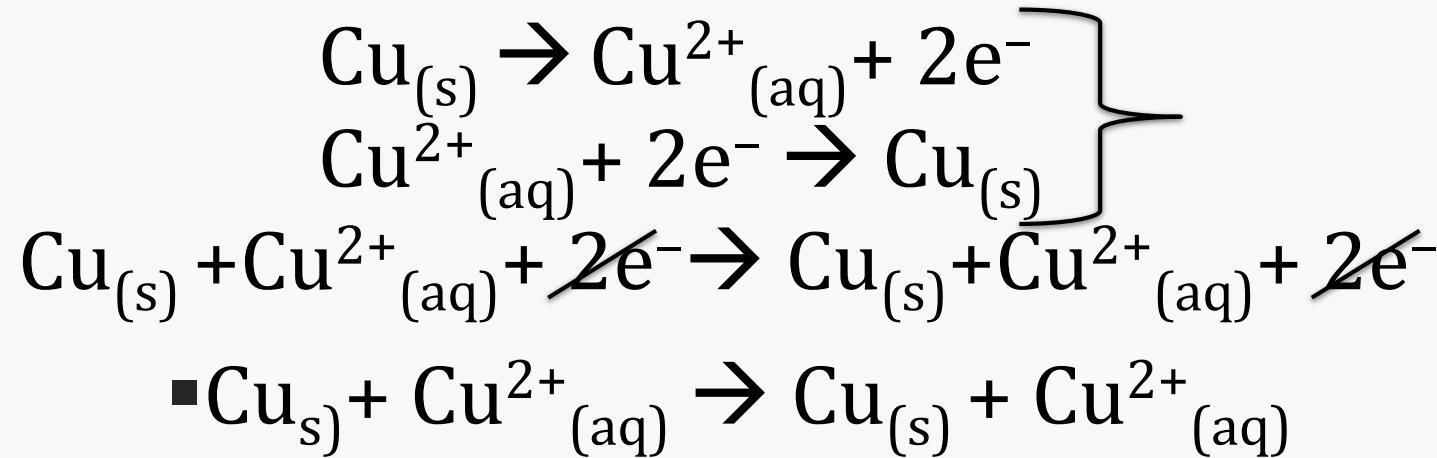
- But conversion of Cu solid from the anode to Cu²⁺ ions into the solution occurs most readily because less energy is required for this process. Therefore, OH⁻ and SO₄²⁻ ions will not be discharged.



- ◻ **At cathode:**

- Ions present are Cu²⁺ and H⁺
- Being lower in electrochemical series Cu²⁺ discharge in preference to H⁺.
- $\text{Cu}^{2+}_{(aq)} + 2\text{e}^- \rightarrow \text{Cu}_{(s)}$
 - ❖ Solid copper is deposited at the cathode

□ *Overall discharge equation:* Combine the equations at anode & cathode



□ Conclusion

- As Cu^{2+} ions from the solution discharge, Cu solid from anode also pass into the solution and form Cu^{2+} ions.
- Therefore, the total concentration of Cu^{2+} and SO_4^{2-} remains constant.
- Thus, blue colour of the solution will be retained.

ii) Mercury used as electrode for the electrolysis of an electrolyte containing sodium ions

- ❖ When an electrolyte containing sodium ions is electrolyzed using platinum cathode, H^+ is discharged in accordance with order of the electrochemical series.
 - Therefore at cathode hydrogen gas will be produced.
- ❖ When mercury cathode is used, there is a possibility of discharging sodium to form **sodium amalgam** with mercury.

Example

- Electrolysis of sodium chloride using mercury electrodes (especially cathode).

◻ Dissociation:

- $\text{NaCl}_{(\text{s})} \rightarrow \text{Na}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$
- $\text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{H}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$

◻ At anode:

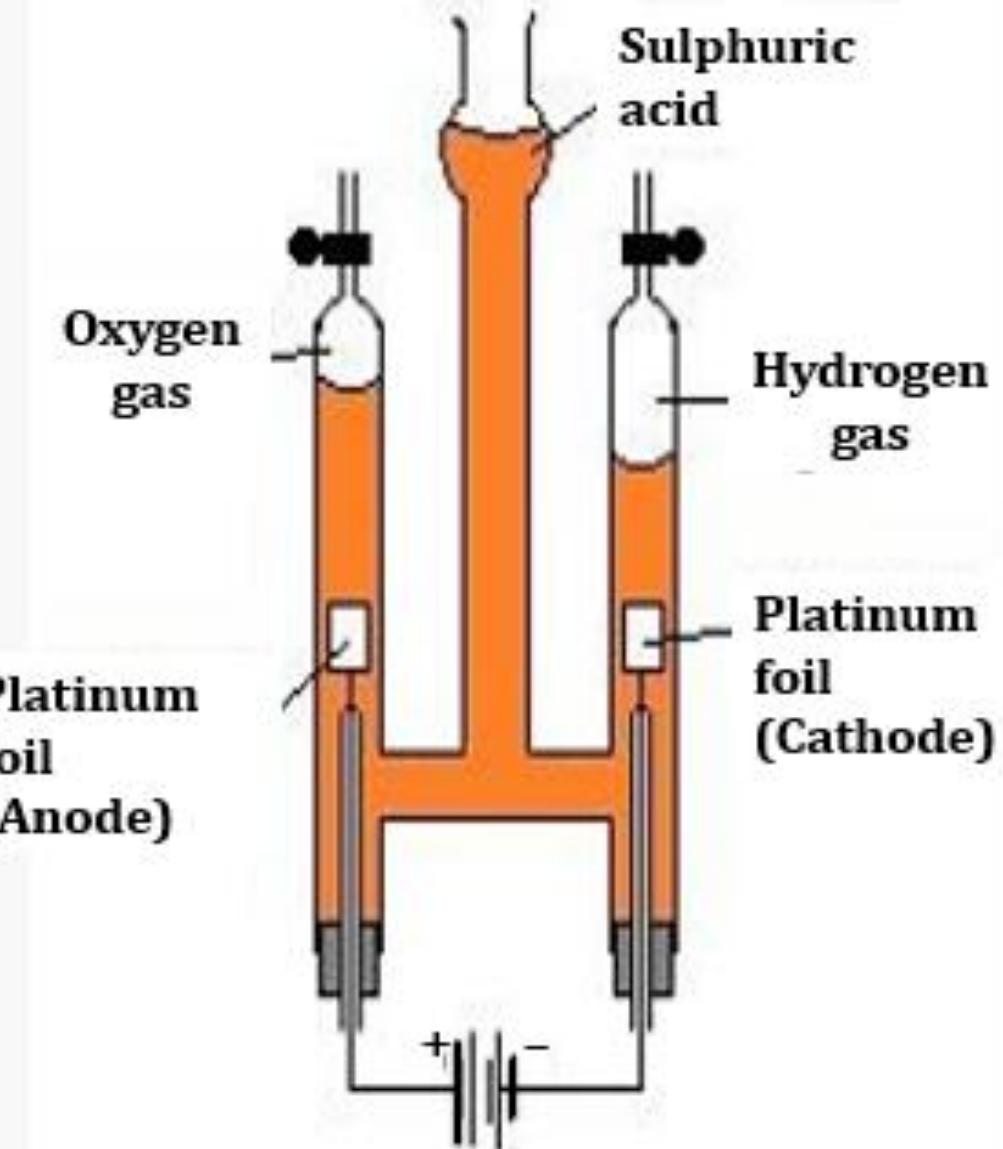
- Ions present are Cl^- and OH^-
- But Cl^- is in higher concentration than OH^- in the electrolyte, so it is preferentially selected for discharge.
- $2\text{Cl}^-_{(\text{aq})} \rightarrow \text{Cl}_2_{(\text{g})} + 2\text{e}^-$
- Chlorine gas is produced at anode

At cathode:

- Ions present are Na^+ and H^+ .
- Na^+ will form sodium amalgam with the mercury at cathode since this process requires less energy than discharge of H^+ .
- Therefore **sodium amalgam** is the product.

Electrolysis of sulphuric acid

- It is conducted in electrolytic cell known as Hoffman's voltammeter.
- Both electrodes used are platinum foils.
- It is also known as *electrolysis of water*.



□ Dissociation of the ions:

- $\text{H}_2\text{SO}_{4(\text{s})} \rightarrow \text{H}^{\text{(aq)}} + \text{SO}_{4}^{2-} \text{(aq)}$
- $\text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{H}^{\text{(aq)}} + \text{OH}^{-} \text{(aq)}$

□ At anode

- Ions present are SO_4^{2-} and OH^{-}
- But OH^{-} , being lower in electrochemical series, is discharged in preference to SO_4^{2-} , in spite of the high concentration of the later.
- $4\text{OH}^{-} \text{(aq)} \rightarrow 2\text{H}_2\text{O}_{(\text{l})} + \text{O}_{2(\text{g})} + 4\text{e}^{-}$
- ❖ Oxygen gas is produced at anode

□ At cathode:

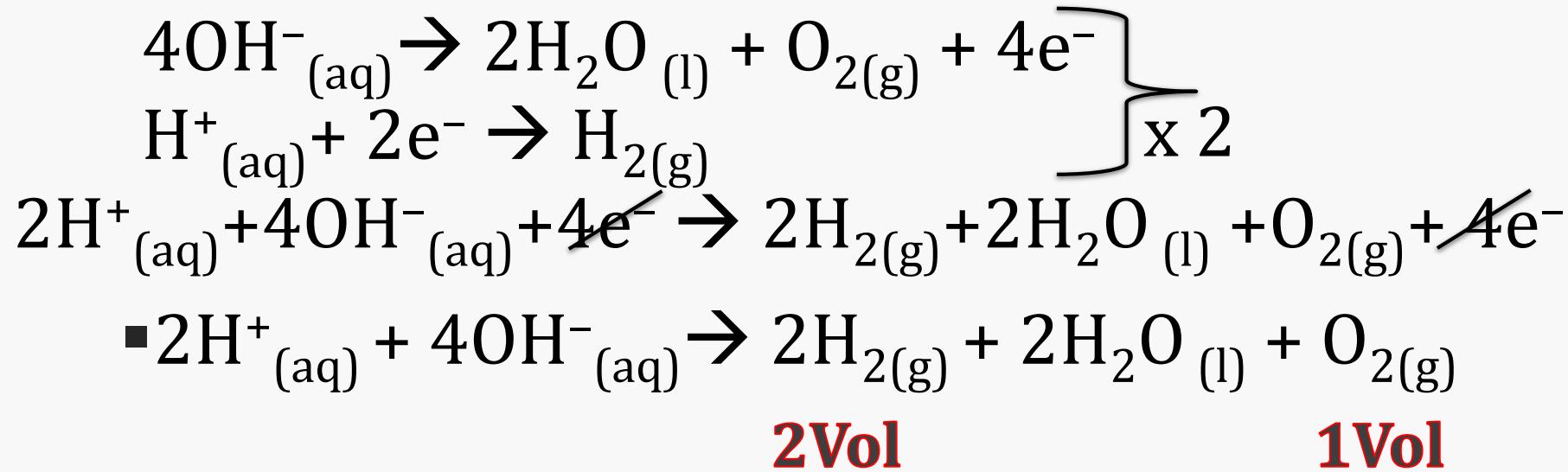
◦ Only H⁺ ions are present at the cathode.

◦ Thus, H⁺ will discharge .



❖ Hydrogen gas is produced at the cathode

□ Overall discharge equation: Combine the equations at anode & cathode



Conclusion:

- Discharge of OH^- at cathode disturbs the ionic equilibrium of water, hence more water ionizes to compensate the loss of OH^- ions in the solution. Excess H^+ produced increases acidity.
- Discharge of H^+ ions at anode lowers the acidity.
- Therefore the total acidity of the solution remains constant.
- ❖ This is why it is called electrolysis of water.

6.3 Laws of electrolysis

- The quantity (number of moles) of a substance produced at an electrode depends on three factors include
 - i. Magnitude of ***current*** passing through the electrolyte per unit time
 - ii. ***Time taken*** to pass the steady electric current through the electrolyte
 - iii. The ***charge*** on an ion
 - ❖ The first and second factor can be combined as quantity of electricity(charge),Q.

$$Q = It$$

The SI unit of quantity of electricity(charge) is coulombs (C)

6.3.2 Faraday's first law of electrolysis

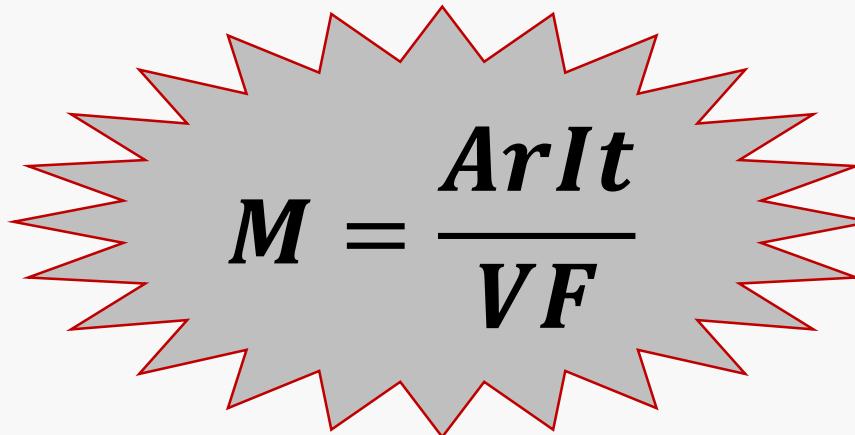
- The law state that “The mass liberated during electrolysis is directly proportional to the quantity of electricity passing through the electrolyte”
- Mathematically $M \propto Q$
- Remove proportionality sign

$$M = kQ$$

- But the proportionality constant, $k = Z$ so we have:

$$M = ZQ$$

- Also $Z = \frac{Ar}{VF}$ and $Q = It$
- Thus the formula becomes:



$$M = \frac{ArIt}{VF}$$

Key

- M = Mass of a substance deposited
- Ar = Atomic mass of the metal deposited
- I = Electric current
- t = time
- Q = The charge (quantity of electricity)
- F = Faraday's constant
- V = Valence
- Z = Electrochemical equivalent.

Important terms to remember

- **A coulomb(C):** is the quantity of electric charge passes through at a given point in a circuit when a current of 1 ampere flows for 1 second.
- **Faraday constant(F):** is the amount of electric charge carried by one mole of electrons($1F = 96500C$).
- **Electrochemical Equivalent (Z):** is the mass of the substance liberated when 1couloumb of electricity is passed through an electrolyte. Its SI Unit is g/c

Example 1:

1) A current of 3.2A is passed through a solution of copper (II) sulphate for 30 minutes. What is the mass of copper deposited at the cathode ($\text{Cu} = 63.5$, $\text{S} = 32$, $\text{O} = 16$)

◦ **Data given:**

*Current, $I = 3.2 \text{ A}$

*Time, $t = 30 \text{ min} = 1800 \text{ sec}$

*Valency, $V = 2$

*Relative atomic mass, $\text{Ar} = 63.5$

*Faraday constant, $F = 96500\text{C}$

**Required:* Mass of copper deposited, M

Solution

◦ From the formula: $M = \frac{ArIt}{VF}$

$$M = \frac{63.5 \times 3.2 \times 1800}{2 \times 96500}$$

$$M = 1.9 \text{ g}$$

❖ Mass of copper deposited at the cathode is 1.9g

Example 2

- How long should a 5A be passed through a molten silver chloride in order to deposit 3.24 Kg of silver at the cathode ($Z = 1.118 \times 10^{-3}$ g/c, Ag = 108)

Data given:

- * Current, $I = 5\text{A}$
- * Mass of silver deposited, $M = 3.24 \text{ kg} = 3240 \text{ g}$
- * Electrochemical equivalent, $Z = 1.118 \times 10^{-3} \text{ g/c}$
- * Time, $t = ?$
- Given that: $1\text{F} = 96500 \text{ C}$

Solution

- From the formula: $M = \frac{ArIt}{VF}$
- But since $Z = \frac{Ar}{VF}$ so, the formula becomes $M = ZIt$
- Making “t” the subject we have: $t = \frac{M}{ZI}$
$$t = \frac{3240}{1.118 \times 10^{-3} \times 5}$$
$$t = 579606.44 \text{ s}$$
- ❖ The current was passed for 579606.44 seconds.

3) Find the quantity of electricity required to produce 10g of aluminium by electrolysis($Al = 27$).

◦ **Data given:**

- * Mass of alumunium, $M = 10g$
- * Relative atomic mass of Al, $Ar = 27$
- * Valency, $V = 3$
- Required : Quantity of electricity, Q

Solution

- From the formula: $M = \frac{ArQ}{VF}$
- Making “Q” the subject we have: $Q = \frac{MVF}{Ar}$
$$Q = \frac{10 \times 3 \times 96500}{27}$$
$$Q = 107222.2 \text{ C}$$
- ❖ The quantity of electrolysis required is 107222.2 C

4) 6 amperes current were allowed to pass through a solution of silver nitrate. At the end of the experiment, 4g of silver deposited at the cathode.

- i) How long was the experiment carried out?
- ii) Find the number of faradays that were passing. ($\text{Ag} = 108$)

◦ **Data given:**

- * Current, $I = 6 \text{ A}$
- * Mass of silver, $M = 4\text{g}$
- * Relative atomic mass, $\text{Ar} = 108$
- * Valency, $V = 1$

Solution

i) Required: Time taken, t

◦ From the formula: $M = \frac{ArIt}{VF}$

◦ Making “ t ” the subject we have: $t = \frac{MVF}{ArI}$

$$t = \frac{4 \times 1 \times 96500}{108 \times 6}$$
$$t = 595.7 \text{ s}$$

❖ Time taken to carry out the experiment is 595.7 seconds

ii) Required: The number of faradays, F

- But, the quantity of electricity, $Q = It$.
- We have: $I = 6 \text{ A}$ and $t = 595.7 \text{ s}$

$$Q = 6 \times 595.7$$

$$Q = 3574.2 \text{ C}$$

- But, recall : $1\text{F} = 96500 \text{ C}$

$$\cancel{x} = 3574.2 \text{ C}$$

$$x = 0.037 \text{ F}$$

❖ There was 0.037 faradays that was passing.

Example 5

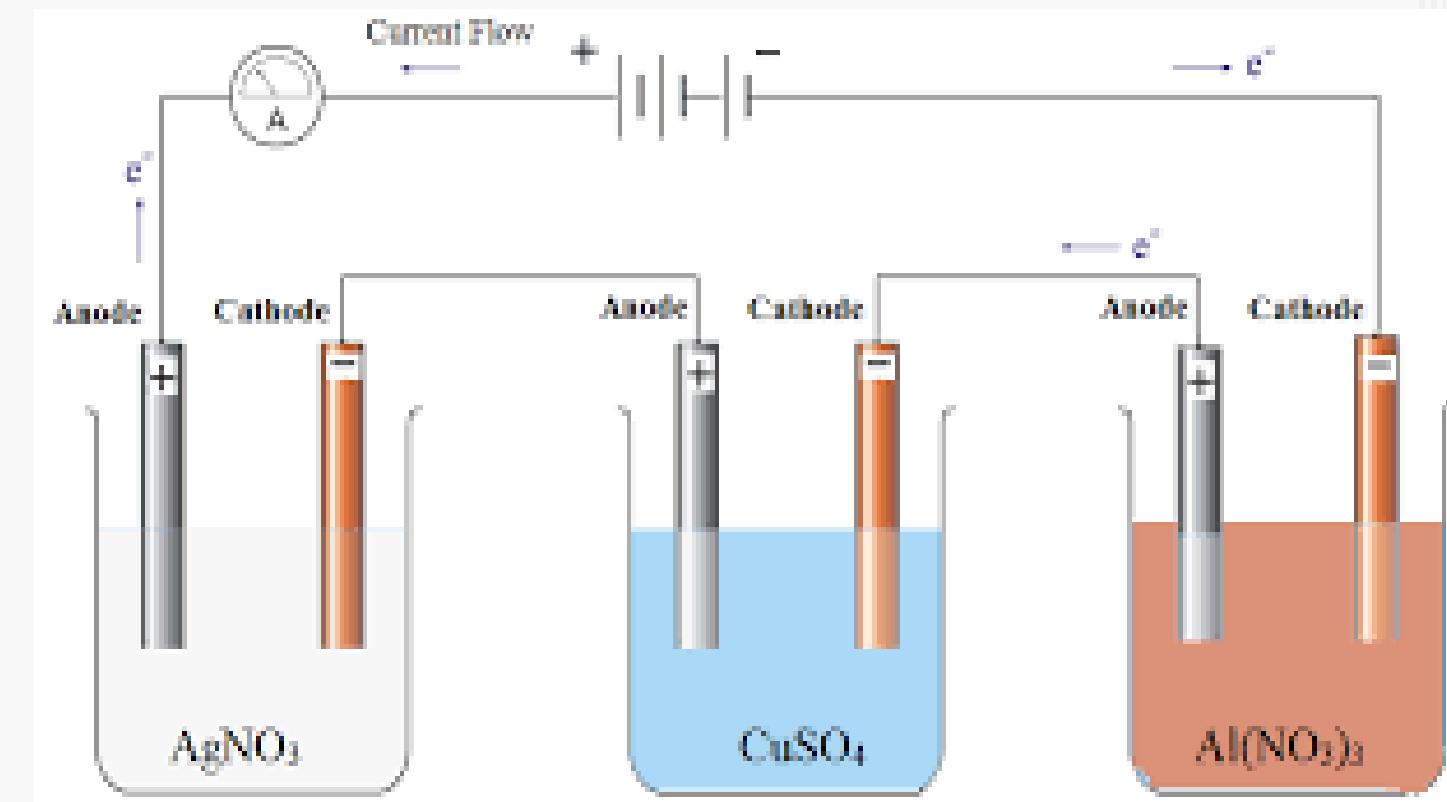
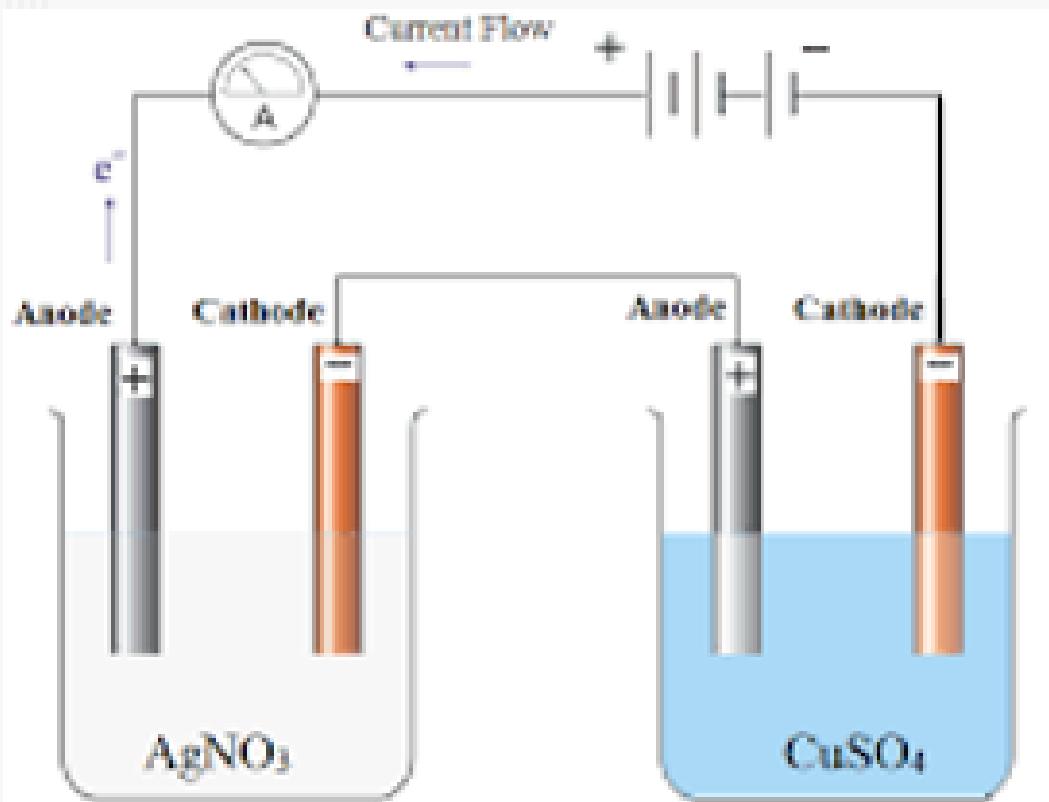
- 0.25 faradays were passing through copper chloride solution using graphite electrodes.
- a) With the aid of discharge equation identify the electrode the following processes occur
- (i) Oxidation (ii) Reduction
- b) Identify the product at
- (i) Anode (ii) Cathode
- a) Calculate
- (i) Mass of a substance deposited at the cathode
 - (ii) Volume of a gas produced at the anode at s.t.p

Try this

1. How many faradays are enough to produce 0.5mole of aluminium by electrolysis?
2. During extraction of sodium by electrolysis of rock salt 6.72 litres of chlorine were produced. Calculate
 - (a) Number of faradays that were passing
 - (b) Mass of sodium produced.

Faraday's second law of electrolysis

- Consider the diagrams below



- The Faraday's second law state that;

"When the same Charge (quantity of electricity) passes through Different electrolyte during electrolysis, the mass liberated is directly proportional to the chemical equivalent"

- Mathematically:

$$M \propto E$$

- To remove the proportionality sign(\propto) we put constant, K

- Making K the subject we get

$$K = \frac{M}{E}$$

- For the First Electrolyte

$$K_1 = \frac{M_1}{E_1}$$

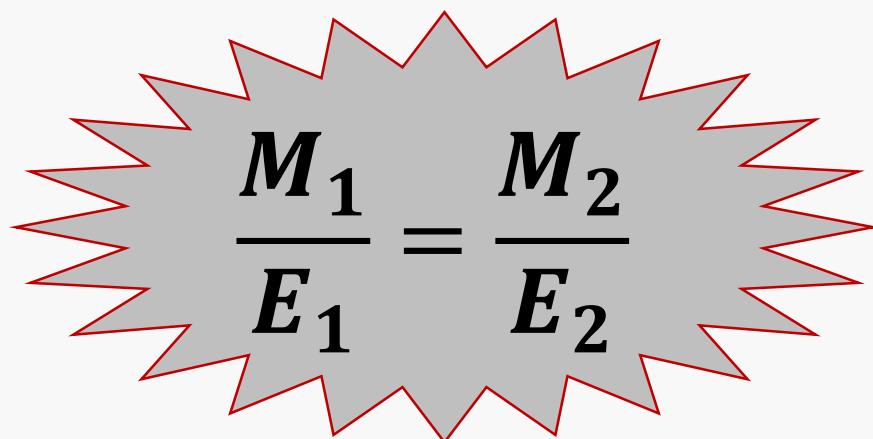
- For the Second Electrolyte

$$K_2 = \frac{M_2}{E_2}$$

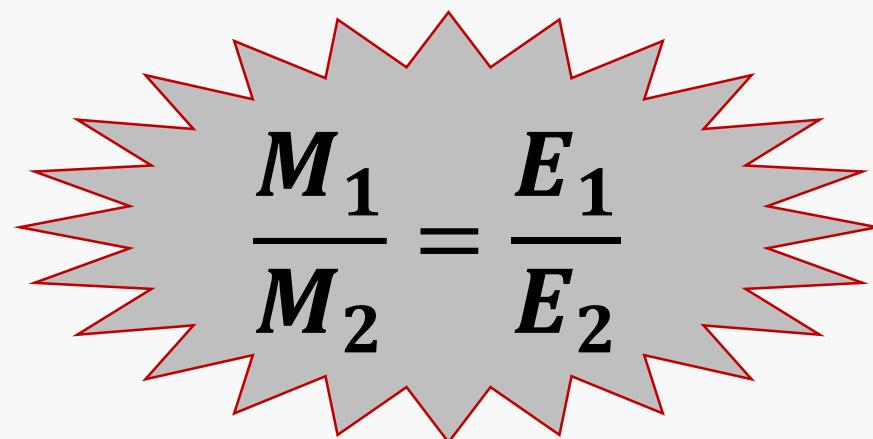
- Since the same quantity of electricity is passing in all the electrolytes, hence

$$K_1 = K_2$$

- Therefore, the formula will be;


$$\frac{M_1}{E_1} = \frac{M_2}{E_2}$$

OR


$$\frac{M_1}{M_2} = \frac{E_1}{E_2}$$

◦ From

$$M = KE$$

But $E = \frac{Ar}{V}$

$$M = \frac{KAr}{V}$$

Divide by Ar both sides we get

$$\frac{M}{Ar} = \frac{K}{V}$$

◦ But $\frac{M}{Ar} = n$, hence

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

◦ This means

$$n \propto \frac{1}{V}$$

◦ This shows that the number of moles is inversely proportional to the charge of the ions.

- From the above relation, second faraday law can also be states as follows;

“When the same quantity of electric current passes through different electrolyte during electrolysis, the relative number of moles of elements deposited are inversely proportional to the number of charges on the ions”

- For first electrolyte

$$n_1 \propto \frac{1}{V_1}$$

$$n_1 = \frac{K_1}{V_1}$$

$$V_1 n_1 = K_1$$

- For Second electrolyte

$$n_2 \propto \frac{1}{V_2}$$

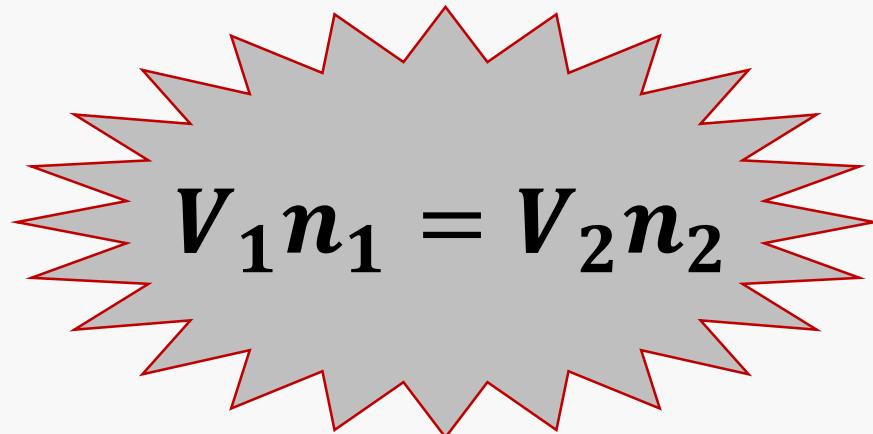
$$n_2 = \frac{K_2}{V_2}$$

$$V_2 n_2 = K_2$$

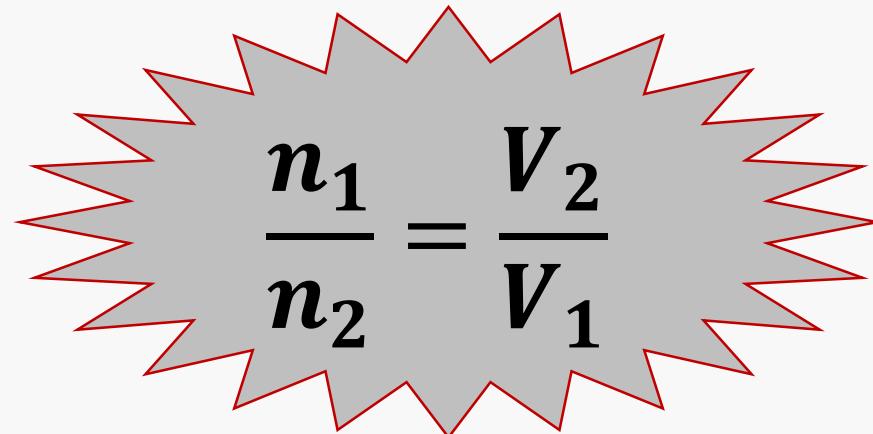
- Since the same quantity of electricity is passing in all the electrolytes, hence

$$K_1 = K_2$$

- Therefore, the formula will be;


$$V_1 n_1 = V_2 n_2$$

OR


$$\frac{n_1}{n_2} = \frac{V_2}{V_1}$$

Where:

- M = the mass liberated during electrolysis
- E = Chemical Equivalent
- Ar =Relative Atomic Mass
- V = Valence (number of charge)
- n = Number of moles

Relationship between electrochemical equivalent(Z) and chemical equivalent(E)

- From the Electrochemical equivalent

$$Z = \frac{Ar}{VF} \text{ ----- (i)}$$

- But also chemical equivalent

$$E = \frac{Ar}{V} \text{ ----- (ii)}$$

- Substitute equation (ii) into equation (i)

$$Z = \frac{E}{F}$$

- ❖ This formula shows the relationship between electrochemical equivalent and chemical equivalent

Example 1

- A current flowing through an electrolytic cell deposits 0.25 g of silver at a certain time interval. How many grams of copper would be deposited from a solution of Copper (II) sulphate in the same time interval using the same current? (Ag=108 Cu=64)

Data given

- Mass of silver $M_1 = 0.25\text{g}$
- Chemical equivalent of silver, $E_1 = \frac{Ar}{V} = \frac{108}{1} = 108$
- Chemical equivalent of silver, $E_2 = \frac{Ar}{V} = \frac{64}{2} = 32$
- Required: Mass of Copper, M_2

◦ From the formula

$$\frac{M_1}{E_1} = \frac{M_2}{E_2}$$

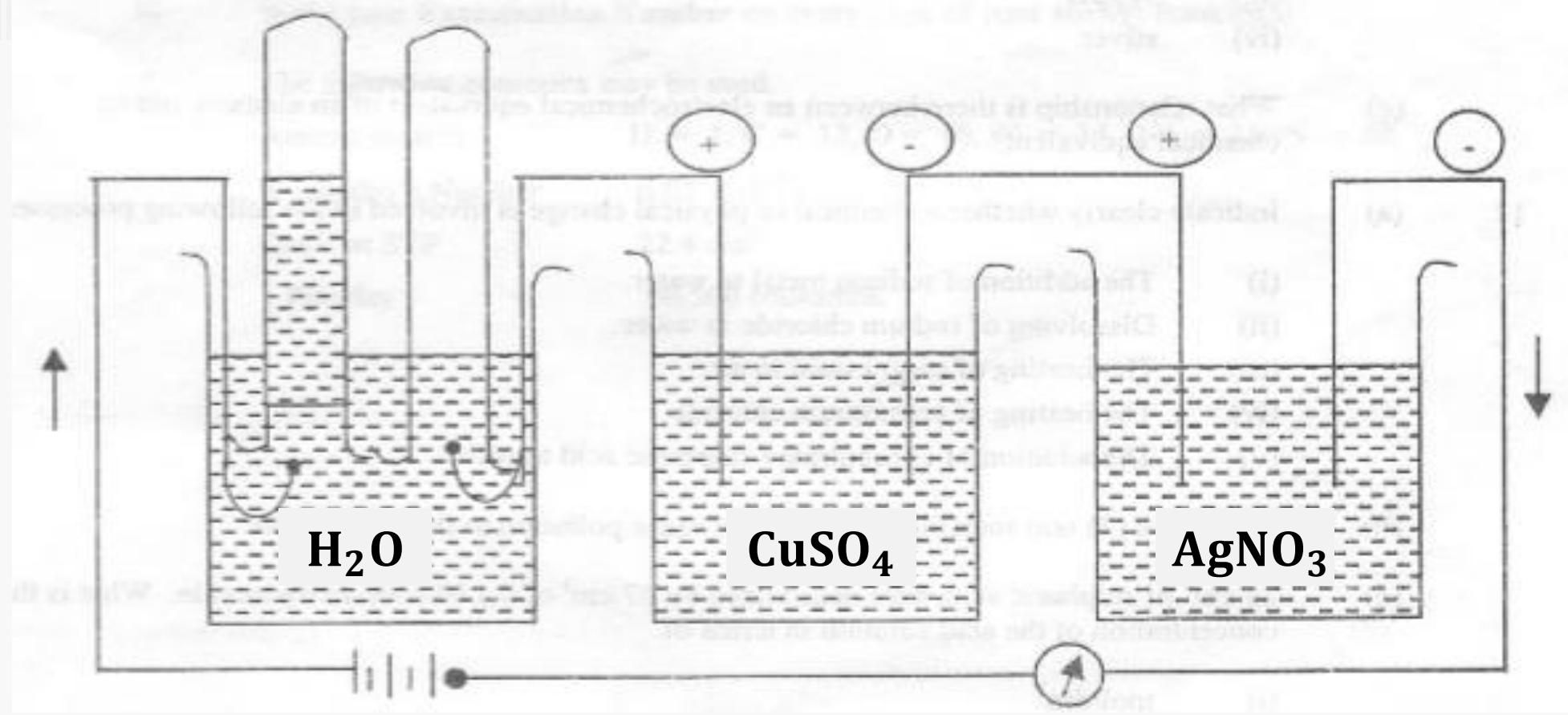
$$\frac{0.25}{108} = \frac{M_2}{32}$$

$$M_2 = 0.074g$$

❖ Mass of copper is $0.074g$

Try this

1. The same current was passing through solution of same concentration of AgNO_3 and CuSO_4 liberated 0.23g of Ag in the first electrolyte. Determine the mass of Cu produced in the same experiment. (Chemical equivalent of Cu = 32 and Ag = 108)
2. The figure below represents an experiment in which Faraday second law was illustrated by connecting in series three cells containing water to which a very little amount of dilute sulphuric acid was added, copper sulphate solution and silver nitrate solution. A current of 1A was passed through a solution for two seconds.



- The volumes and hence weights of hydrogen and oxygen liberated were calculated. The weight of copper and silver formed by the electrolysis of copper sulphate and silver nitrate solutions was also measured.

◦The result of this experiment are tabulated below

Element	Current, I (A)	Time (s)	Mass of element deposited (g)	Quantity of electricity, Q (C)	Electrochemical equivalent, Z
1 Hydrogen	1.00	2	2.0892×10^{-5}		
2 Oxygen	1.00	2	1.658×10^{-4}		
3 Copper	1.00	2	6.587×10^{-4}		
4 Silver	1.00	2	2.236×10^{-3}		

a) Complete the table above by calculating

- Quantity of electricity passed in experiment 1,2,3 and 4.
- Electrochemical equivalent of the elements in experiment 1,2,3 and 4.

b) If the faraday constant is 96500C, calculate the chemical equivalents of

- i) Hydrogen
- ii) Oxygen
- iii) Copper
- iv) Silver

c) What relationship is there between an electrochemical equivalent and it's chemical equivalent.

6.4 Applications of electrolysis

- i) ***Extraction of metals:*** Very reactive elements are extracted by electrolysis of their molten chlorides or oxides, for example Al, Mg, Na, K (Details in extraction of metals)
- ii) ***Purification metals:*** In this process a block of the impure metal is molded and placed at the anode and sheets of extremely pure metal are made at the cathode
- iii) ***Preparation of gases:*** During electrolysis gases are liberated at anode except hydrogen.
- iv) ***Electroplating:*** This is the process of covering the corroding material (metals) by the unreactive metals using electrolysis to prevent corrosion and improve their appearance.

Importance of electroplating

- It improves appearance
- It protect the object from getting rust

MORE PRACTICES

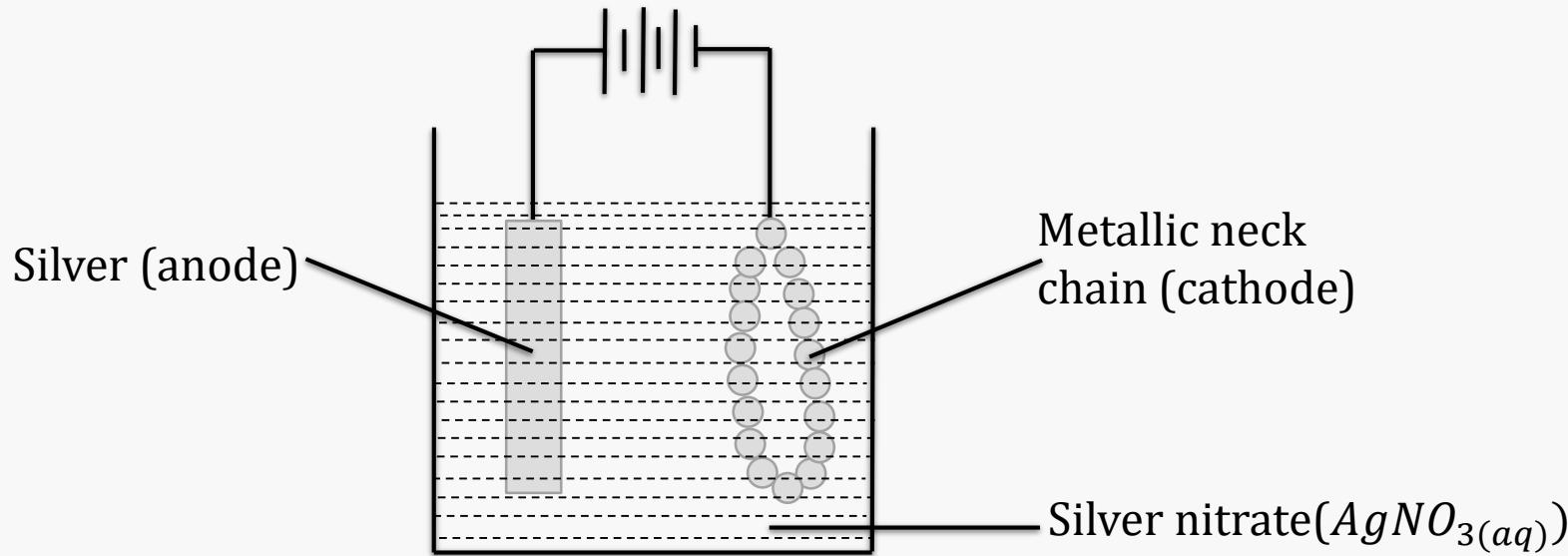
1. 0.2 faradays were passing through dilute sulphuric acid which was electrolyzed by using platinum electrodes.
 - a) Write the chemical equations for the discharging process at the anode and cathode when.
 - b) Calculate the volume of gases produced during this experiment.
2. Briefly explain the nature of the remaining solution and write the chemical equations for the discharging process at the anode and cathode when copper(ii) sulphate is electrolyzed using
 - a) Carbon electrodes
 - b) Copper electrodes

3.(a) 0.02 mole of electrons was passed through a solution of sodium hydroxide using platinum electrodes.

- (i) Give the names of the gases evolved at each electrode.
 - (ii) Write the ionic equations of the reaction taking place at the electrodes.
 - (iii) Calculate the number of moles of each gas produced and the volume which each gas would occupy at S.T.P.
- (b) What mass of copper will be liberated during electrolysis of copper sulphate solution by charge of one faraday?

4. An element X has a relative atomic mass of 88. When a current 0.5A was passed through the fused chloride of X for 32 minutes and 10seconds, 0.44g of X was deposited at the cathode.
- (i) Calculate the number of faradays needed to liberate 1mole of X.
- (ii) Write the formula for the X ion.
- (iii) Write the formula for hydroxide of X.
5. 289500 coulombs were required to deposit one mole of a metallic element Q from it's aqueous salt solution. calculate the valence of Q.

6. The following experiment was used to electroplate a metallic neck chain. Study and answer the questions that follow:



- Explain what happened to the anode and cathode
- Write the ionic equations for the reaction which occurred at the electrodes.
- Calculate mass gained by neck chain if 0.25F was allowed to pass through the electrolyte in this experiment.

7. Assume that you are a chemist in a chemical plant that deals with the production of chlorine gas. You want to produce 100 litres of chlorine gas per hour so that you can reach the company's goal of producing 2400 litres every day. What current of electricity will you allow to flow per hour?
8. How long a current of 5A should be passed through a solution of silver chloride in order to deposit 3.24 g of silver metal at the cathode? Given that, the electrochemical equivalent of silver, $Z_{Ag} = 1.118 \times 10^{-3} ge^{-1}$.

7. An electric current was passed in series through solutions of calcium chloride and copper(ii) sulphate. Carbon electrodes were used in both electrolytes. If 2.5L of chlorine measured at S.T.P were produced.

- a) Write the ionic discharge equations for the reactions taking place at cathodes and anodes both two cases.
- b) What volume of oxygen would also be produced?
- c) What mass of copper was produced?

TOPIC 7

CHEMICAL KINETICS, EQUILIBRIUM AND ENERGETICS

7.1 RATE OF CHEMICAL REACTIONS

Chemical kinetics:

- Is the study of the rate of chemical reaction.
- It involves investigations into how different conditions influence the rate/speed of a chemical reaction and the process through which the chemical reaction takes place.

Rate of reaction:

- Is the measure of the progress of the reaction per unit time.
- Consider the following hypothetical reaction:



- The equation shows that B disappears 3 times as fast as A

- In any chemical reaction the proportions of both reactants and products changes.
- Normally reactants decrease forming products, thus products increase.
- The reaction reaches to completion when at least one of the reactants **depletes**. The reactant which depletes is called *limiting reagent* or *limiting reactant*.
- **Limiting reagent** is the reactant in the chemical reaction that determines the amount of product that is formed.

OR

- **Limiting reagent** is the reactant that is fully consumed when a reaction goes to completion.

- For the reversible reactions the reaction reaches at *equilibrium* and does not go to *completion*.
- The rate of reaction can be determined by measuring the amount of *reactant used up* or *product formed* at a given time.

$$\text{Rate of reaction} = \frac{\text{Reactant used}}{\text{Time}}$$

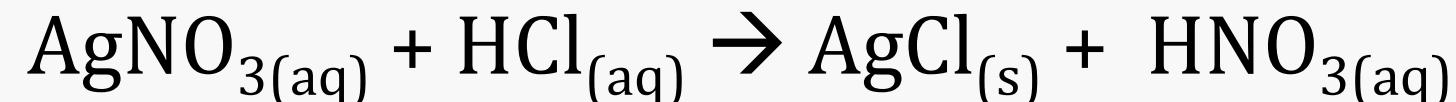
OR

$$\text{Rate of reaction} = \frac{\text{Product formed}}{\text{Time}}$$

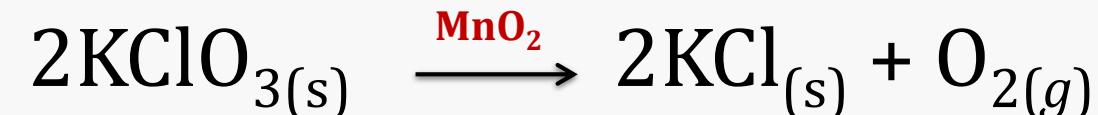
- Depending on the rate of reaction the, the reactions can be grouped into three categories

i) Fast reaction: Is the reaction which occurs within a fraction of second.

Example: Ionic precipitation reactions



ii) Moderate reaction: Is the reaction which occurs within a couple of minutes or hours.



iii) Slow reaction : Is the reaction which takes place gradually and can continue for a number of years. Example Rusting and Radioactive decay

COLLISION THEORY AND ACTIVATION ENERGY

- According to collision theory the reaction occurs due to the ***approach*** and ***collision*** of reacting particles.
- The reacting particles may be ***ions, atoms or molecules.***
- Only collisions that succeed are the one which produce chemical reaction.
- The successful collision are known as ***effective collisions.***
- The colliding particles must possess ***activation energy*** which results in chemical reaction.

- **Activation energy(E_a):** Is the minimum energy that the reacting species must possess in order to undergo a specified reaction.

NOTE: In many spontaneous exothermic reactions at room temperature, most of the molecules simply rebound from collision without reacting.

- **Activated complex (transition state) :** Is the weak intermediate product which may lead to stable products.

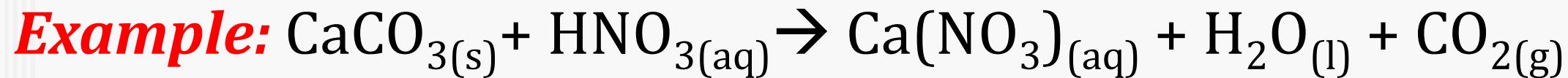
7.2 FACTORS AFFECTING THE RATE OF CHEMICAL REACTION

i) Surface area

- Increases the Surface area tends to increases the rate of chemical reaction
 - This increases more effective collision of reactants particles by increase more number collision of reactant particles per time.
 - The affects only heterogeneous reaction.
- ❖ ***Homogeneous reaction*** is the reaction whose reactants are in the same physical state.



❖ ***Heterogeneous reaction*** is the reaction whose reactants are in different physical state.



ii) Temperature

- The increase in temperature increases the rate of chemical reaction.
- This is because increases in temperature tends to increase kinetic energy, hence more effective collision of reactants particles.
- This supplies the reactant particles with the activation energy.

iii) Concentration

- Increases the concentration tends to increases the rate of chemical reaction.
- This is because number of particles per volume increases, hence more effective collision of reactants particles.

iv) Pressure

- When pressure increase the rate of chemical reaction increases.
- From Boyle's law pressure and volume are inversely related, so increase in pressure tend to increase rate of chemical reaction by decrease volume of reactants result greater chance of collision.

v) Light

- Light increase rate of reaction some reaction like photosynthesis process in plant, present of light favour production of carbohydrate and when absent no photosynthesis process happen

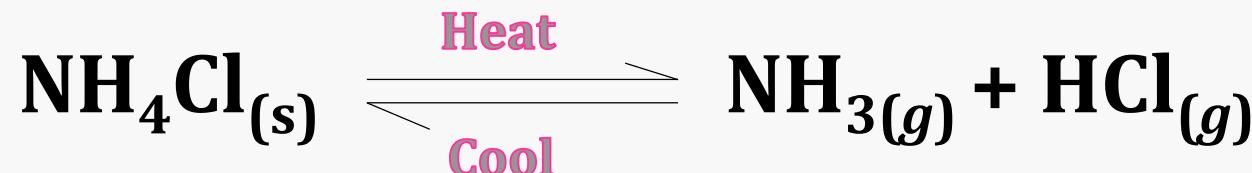
vi) Catalyst

- Catalyst increases the rate of chemical reaction by lowering the activation energy required to start a reaction.
- Thus when the catalyst is added the reactant particles attains activation energy faster.

Reversible and irreversible reactions

- **Reversible reaction:** Is the reaction which can proceed in either direction depending on the condition in which it is carried out.
- This kind of reaction can never reach to the completion but it can only reach to the *equilibrium*.
- This type of reaction is indicated by double half curled arrows showing that it can proceed to the product or reactant side.

Example:



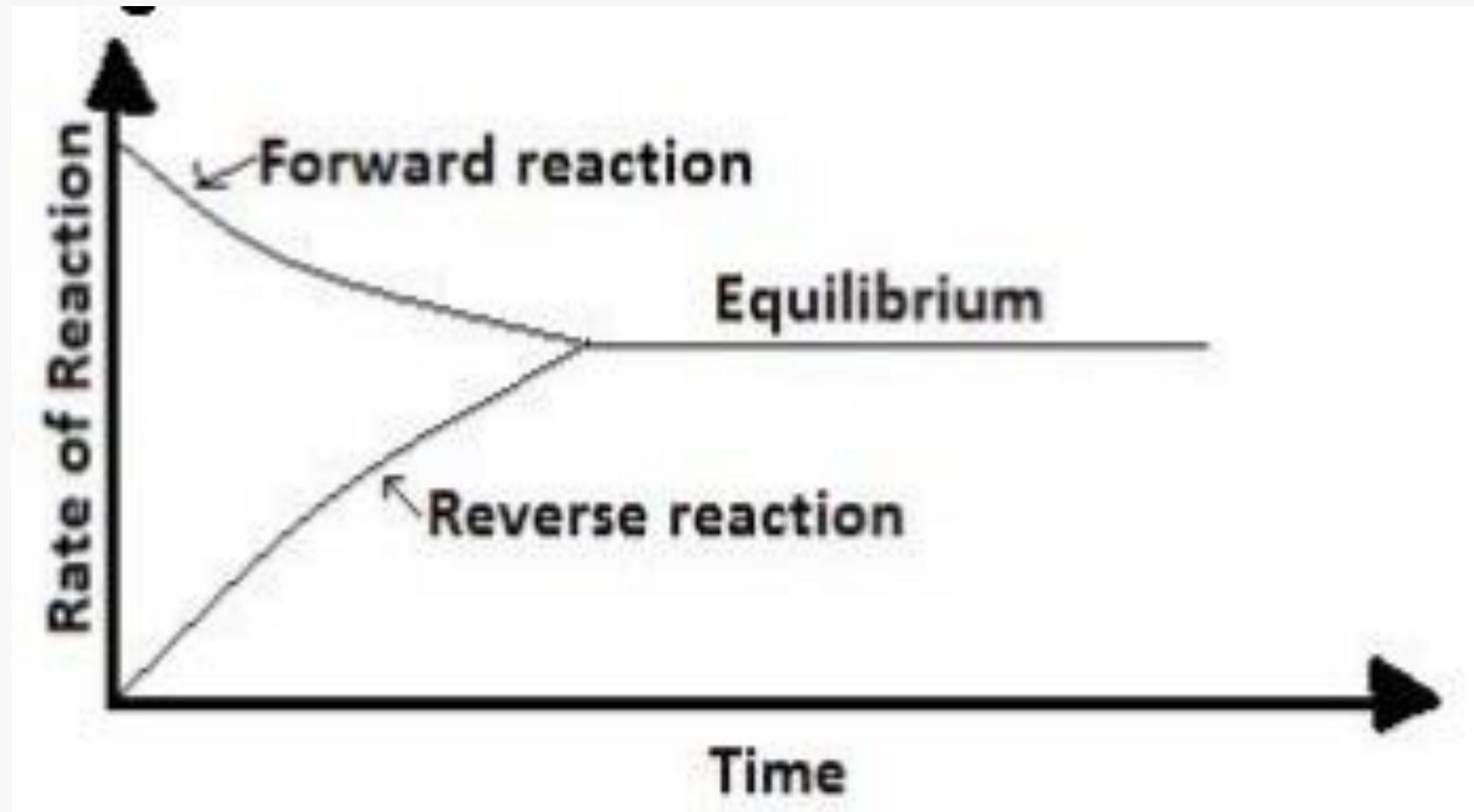
- **Irreversible reaction:** is the kind of chemical reaction which proceeds only in forward direction.
- This kind of reaction reaches to the completion when at least one of the reactants is depleted.
- This type of reaction is indicated by a single full curled arrow showing that it proceeds in only one direction.
- Example:



7.3 Chemical equilibrium

- Chemical equilibrium is the relative stability of the reactant and the products.
- Chemical equilibrium occurs in a ***closed system***.
- **Closed system:** is the situation of reversible reaction where no product or reactant escape to the outside environment.
- **Equilibrium reaction** is the reaction whereby the rate of forward reaction and backward reaction are equal.

- **Dynamic equilibrium** is the situation where backward and forward reaction proceed at the same rate at the same time.



QUIZ

1. Define the following
 - a) Rate of chemical reaction
 - b) Activation energy
 - c) Chemical kinetics
2. Mention any four factors affecting the rate of chemical reaction.
3. Zinc granules were allowed to react with dilute hydrochloric acid to form hydrogen gas. 5 litres of hydrogen were produced for 30 minutes
 - a. Calculate the rate of chemical reaction (ans: 0.0028 litre/sec)
 - b. Briefly explain one factor that should be varied so as to increase the rate of chemical reaction

Le Chatelier's Principle

- It states that:

“If stress is applied to a system at equilibrium, the system readjusts, to counteract the effect of the stress”

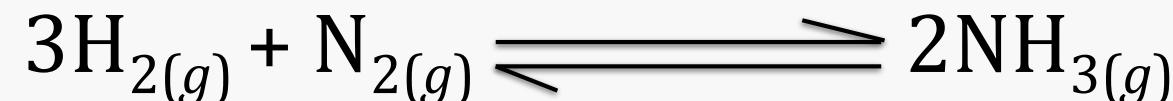
Factors affect position of equilibrium

- i) Concentration
- ii) Pressure
- iii) Temperature

i) Concentration

- Concentration does not affect solids at all.
- The reaction always shift (proceeds) from the side of ***high concentration to low concentration.***
- Thus, if the concentration of the reactants increases the position of equilibrium shifts forwards (to the products side).

Example Consider the reaction below:



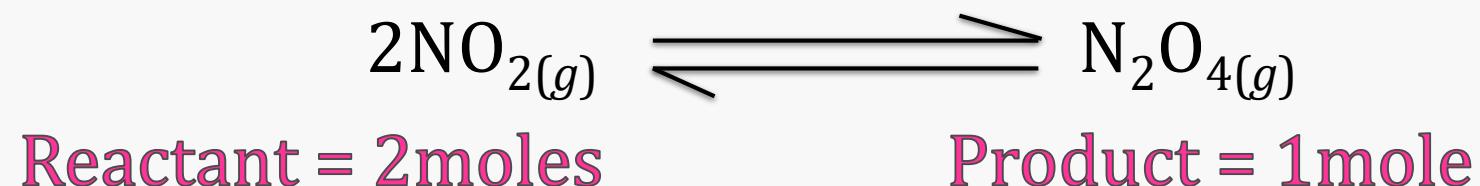
➤ If the concentration of H_2 or N_2 increase the position of equilibrium shifts forwards(to the product side).

- If the concentration of NH_3 increases the position of equilibrium shifts backwards(to the reactant side)

ii) Pressure

- Pressure affects only gaseous components of the reaction, because they are compressible.
- An increase in pressure favours the a reaction which forms fewer moles and vice versa.
- If the number of moles are equal pressure will have no effect.

Example Consider the reaction below:



- If pressure is increased the position of equilibrium will shift forward.
- If pressure is lowered the position of equilibrium will shift backward.

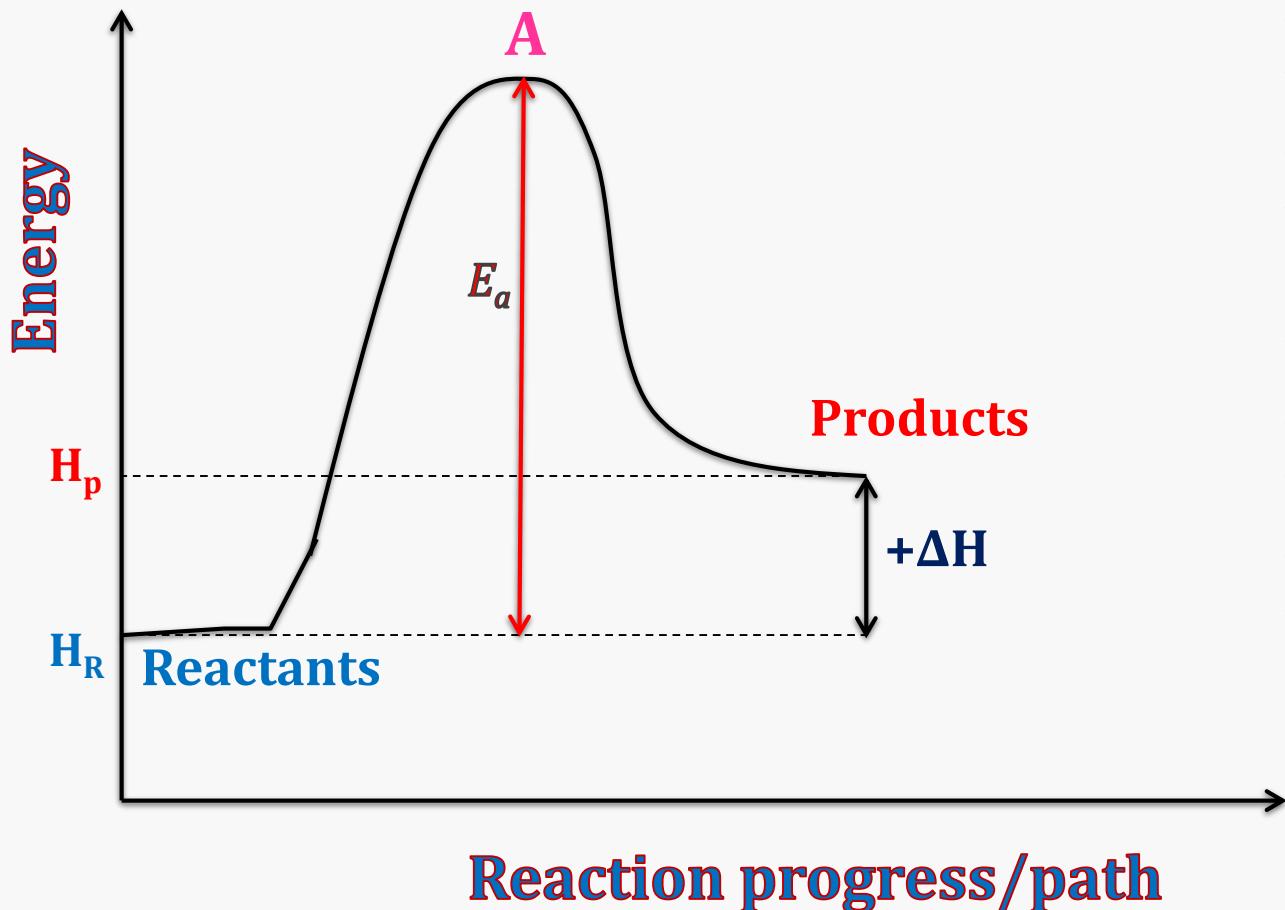
iii) Temperature

- If the temperature increases the position of equilibrium shifts to the side where heat energy is absorbed.
- Therefore the effect of temperature to the position of equilibrium depends on the type of reaction, which are;
 - a) Endothermic reaction
 - b) Exothermic reaction

a) Endothermic reaction is the reaction which absorbs heat energy from the surrounding.

- Increase in temperature favour endothermic reaction.
- It's heat change is positive, it is denoted by $+ΔH$ kJ.
- ❖ Example of endothermic reactions:
 - Dissolving hydrated sodium thiosulphate in water
 - Dissolving ammonium nitrate in water
 - Mixing water and potassium chloride
- ❑ When temperature is increased in endothermic reaction, the position of equilibrium will shift forwards(to the products)

Phase Diagram Of Endothermic Reaction



Key:

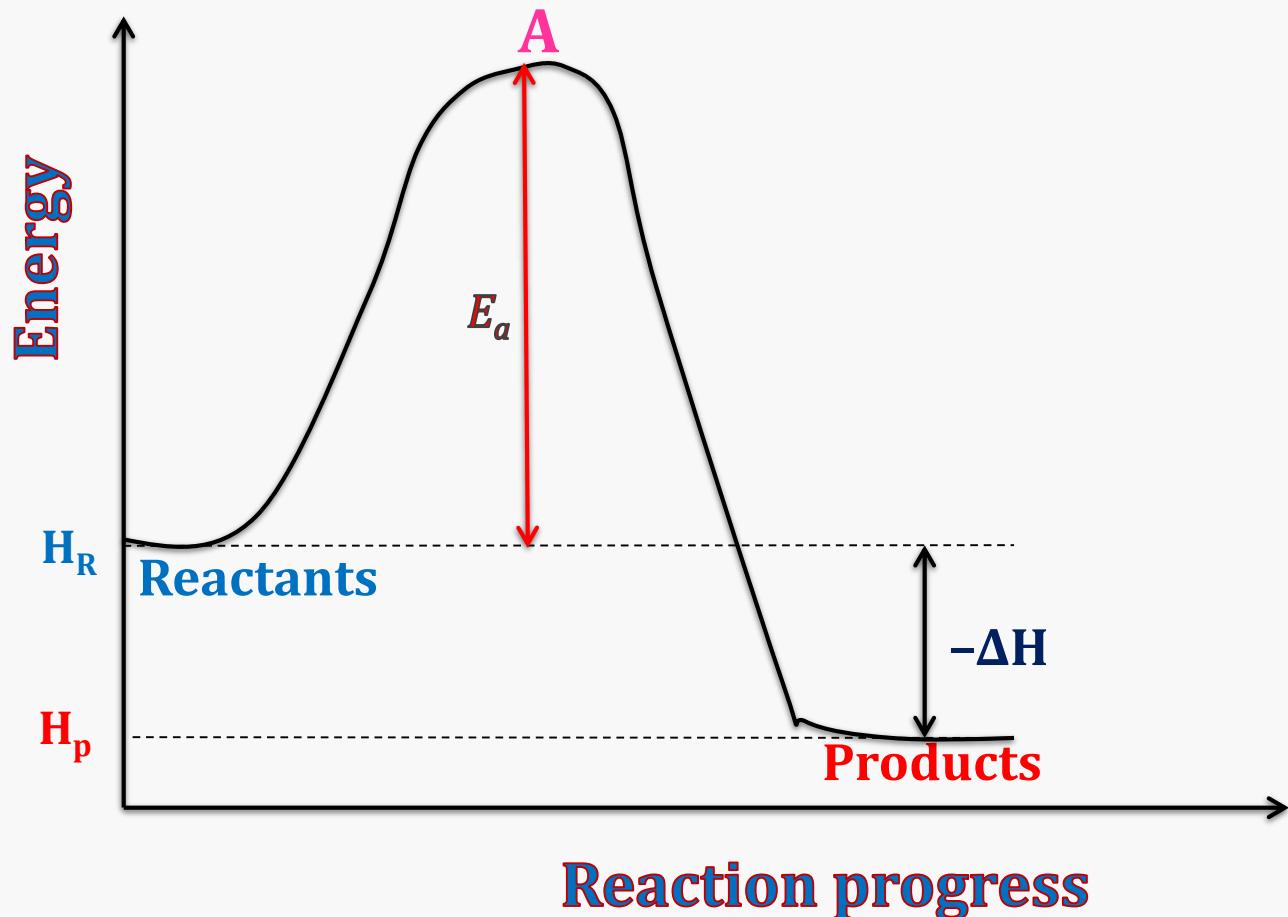
- A = Activated complex
- E_a = Activation energy
- H_R = Heat energy of the reactant
- H_p = Heat energy of the products
- $ΔH$ = Enthalpy (heat) change

$$\diamond H_p - H_R = +\Delta H$$

b) Exothermic reaction is the reaction release heat energy to the surrounding

- Increase in temperature does not favour exothermic reaction.
- It's heat change is negative, it is denoted by $-\Delta H$ kJ.
- ❖ Example of exothermic reactions:
 - Mixing of water and strong acids like Concentrated HCl.
 - Mixing of water and strong base like NaOH.
 - Combustion reactions.
- ❑ When temperature is increased in exothermic reaction, the position of equilibrium will shift backwards(to the reactants)

Phase Diagram Of Exothermic Reaction



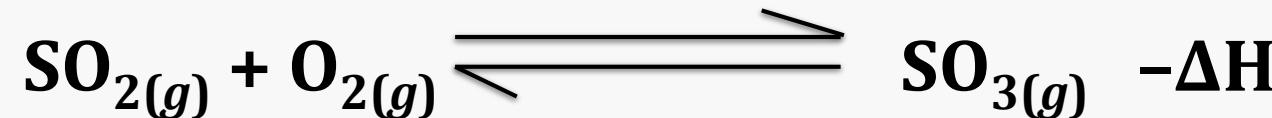
Key:

- A = Activated complex
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- H_p = Heat energy of the products
- ΔH = Enthalpy (heat) change

$$\diamond H_p - H_R = -\Delta H$$

Try this

1) The following is the reaction which takes place in one of the stages of manufacturing of sulphuric acid:



◦ Briefly explain what will happen to the position of equilibrium if:-

- a) Pressure increases
- b) Temperature is increased
- c) All SO_2 is removed from the system
- d) O_2 is added to the system
- e) More SO_3 is added to the system
- f) Volume of the container is increased.

2) Hamis was preparing carbon dioxide using marble chips and 0.1M HCl. He find that the reaction was very slow. Explain briefly how will you help Hamis to increase the rate of production of CO_2 .

3) In an experiment to determining the rate of reaction, the following volumes of gas in the syringe were recorded in every 10 seconds.

Time (sec)	0	10	20	30	40	50	60	70	80
Volume (cm^3)	0	18	30	40	48	53	57	58	58

a) Plot a graph of the volume against time.

b) Why was there no more increase in volume between 70th and 80th sec?

Equilibrium of reactions of industrial importance

- The industrial plants are designed in such a way they operate at optimum production level and at least possible cost.
- This involves choosing the right conditions for obtaining reasonable yields.
- The knowledge of dynamic equilibrium helps in the choice of the appropriate condition.
- The industrial processes which applies this knowledge are
 - a) Haber process
 - b) The contact process

a)Haber process

- This is the process by which ammonia gas is manufactured from hydrogen and nitrogen.
- The chemical equation for the reaction is as follows;



i) Effect of temperature

- This is exothermic reaction, hence the increase in temperature will lower the yield of ammonia as equilibrium shift backwards.
- However at room temperature the reaction is very slow.
- Therefore temperature of about 500°C is used.

ii) Effect of pressure

- The increase in pressure favours forwards reaction, hence high yield of ammonia.
- At temperature 200°C and pressure above 750atm, there is 100% conversion of reactants to product but it is difficult and expensive.
- Thus, a lower pressure of 200atm and 500°C is usually used where the rate of production of NH_3 is 10% – 20%.
- These conditions gives an acceptable yield of NH_3 within an acceptable period of time.

b) The contact process

- This is the process used in industrial manufacture of sulphuric acid.
- Conversion of SO_2 to SO_3 which is an equilibrium reaction



i) Effect of pressure

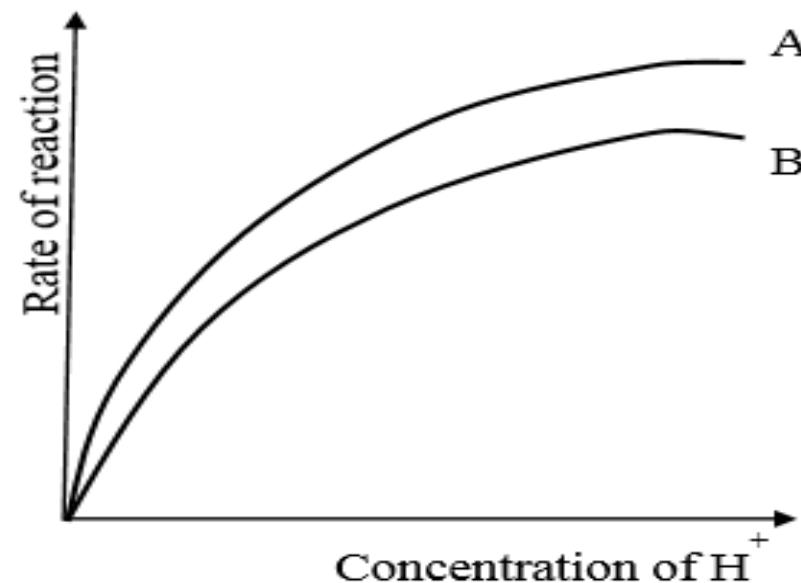
- The increase in pressure favours forwards reaction, hence high yield of Sulphur trioxide.
- But to maintain high pressure is very expensive and difficult.
- Therefore the reaction is maintained at atmospheric pressure(1atm) where by yield of SO_3 is 90%.

ii) Effect of temperature

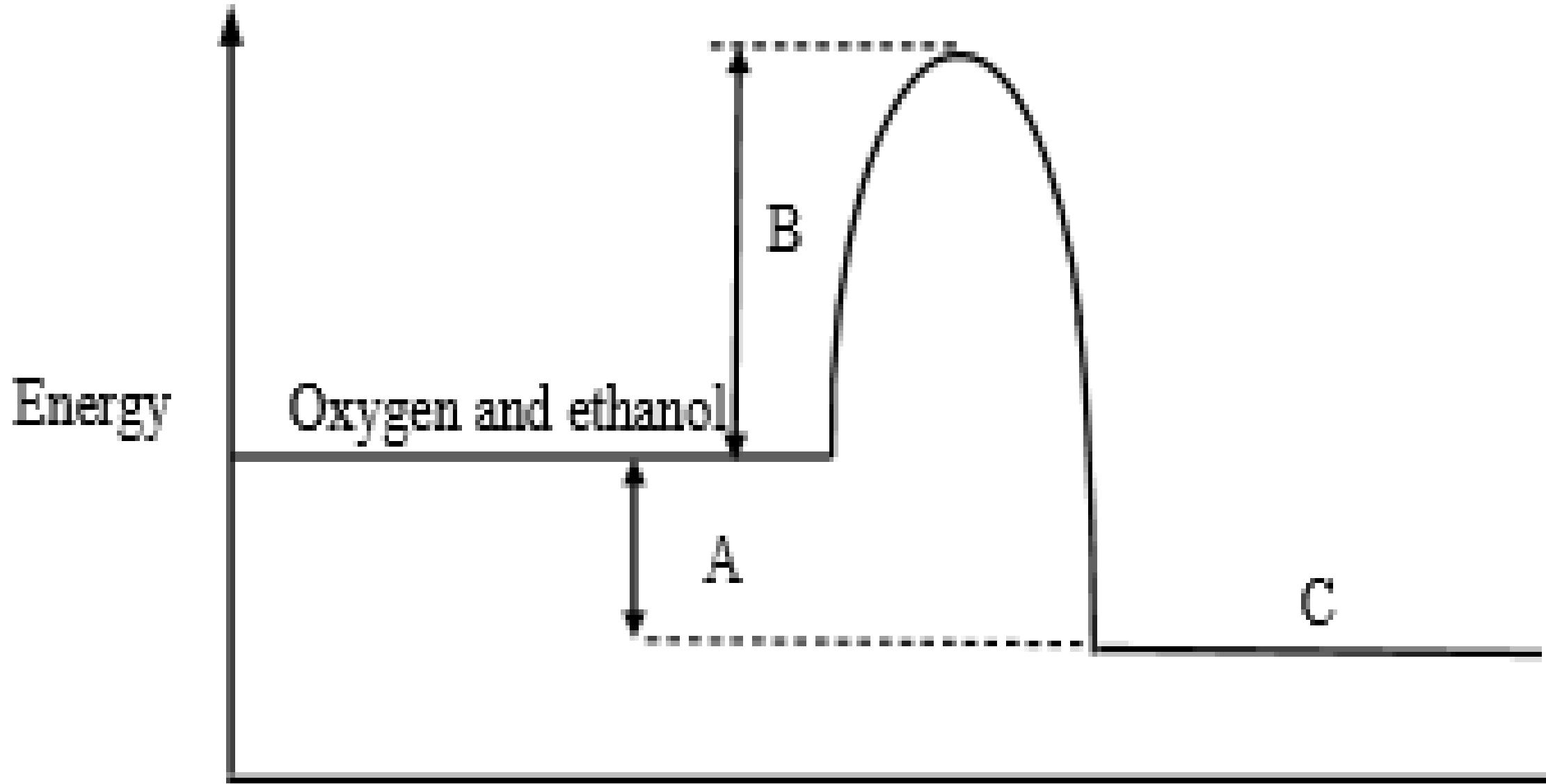
- The reaction is exothermic reaction, hence favoured by low temperature.
- However at low temperature the reaction is very slow, thus the temperature between 400°C and 500°C is used for acceptable rate of reaction(normally 450°C).
- Vanadium penta-oxide or platinum is used as catalyst for this reaction.

More problems

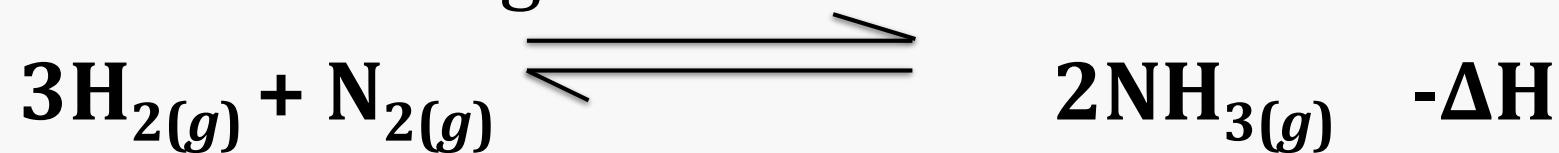
1) Two experiments were carried out using the same mass of magnesium ribbon and the same volume of acids of the same concentration. The acids were 1M hydrochloric acid and 1M ethanoic acid. The results were as shown in the following figure:



- (a) If the experiments were conducted within the same time, is there a difference in volumes of hydrogen gas collected at the same room temperature and pressure? Give reasons for your answer.
- (b) When same mass, volume and concentration of powdered magnesium and ethanoic acid are allowed to react, new graph is formed. Giving reason (s), suggest the position of that graph whether will be above, between or below graphs A and B.
- 2) Oxygen and ethanol react to produce carbon dioxide and water according to the following energy level diagram:



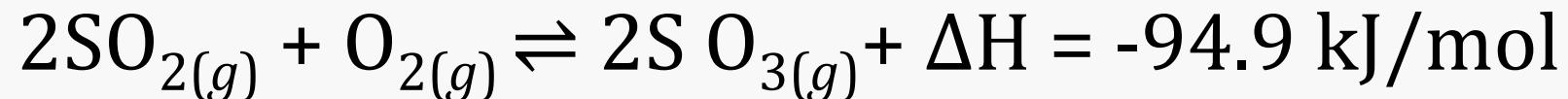
- (i) What is represented by letter A, B and C?
- (ii) What type of reaction is represented by this energy level diagram?
- 3) The following chemical equation represents the reaction to prepare ammonia gas



Explain how will you adjust the following conditions in order to increase the yield of ammonia gas.

- i) Pressure
- ii) Temperature
- iii) Concentration

4. In the industrial preparation of sulphur trioxide, equilibrium is established between sulphur dioxide and oxygen gas as follows:



- (a) (i) Is the forward reaction an endothermic or exothermic process? Give a reason.
- (ii) How would you adjust temperature and pressure to maximize the proportion of the product at equilibrium?
- (b) (i) Why is it unfavourable to work with very high pressure and very low temperature in the Contact process?
- (ii) What catalyst is used to speed up the rate of formation of sulphur trioxide before attaining equilibrium?

PRACTICALS

Possible errors

- i. Operation of the stopwatch
- ii. Over diluting the solutions.
- iii. Misreading the volume (parallax problem or error in counting unmarked graduation marks)
- iv. Using contaminated solutions
- v. Using dirty glass
- vi. Losing the solution while mixing.

Precautions

- i. Be careful and learn how to operate the stopwatch before experiment
- ii. Dilute the solutions careful if you are told to dilute to avoid over diluting it.
- iii. The eye should be kept in level with liquid surface while taking measuring cylinder readings.
- iv. Handle the solution with care to avoid contaminations.
- v. Wash the apparatuses well before experiment.
- vi. Make sure you don't lose even a drop of solution.

TOPIC 8

EXTRACTION OF

METALS

8.1 Occurrence of Metals

- Metals are the elements which form positive ions by losing electron.
- Since metals they form positive ions by losing electrons are also known as electropositive elements.
- The least reactive metals may be found free in nature but the more reactive metals are found in combined state.
- In our country Tanzania there are many different places where different metals are extracted.

	Metal /Ore	Region
1.	Gold	Geita, Rukwa
2.	Tin	Bukoba
3.	Copper	Singida, Rukwa
4.	Diamond	Tabora
5.	Limestone	Dar-es-Salaam, Pemba, Unguja.
6.	Gypsum	Tanga
7.	Salt	Kigoma (Uvinza)

Physical properties of metals

- i) They are Good electrical conductors.
- ii) They are Good heat conductors.
- iii) They are Malleable (ability to bend or to be hammered into other shapes without breaking).
- iv) They are Ductile (ability to drawn/stretched into wire).
- v) They are shiny lustre (shine when polished).
- vi) They are opaque as thin sheet.
- vii) They are Solid at room temperature except Mercury.

8.2 chemical properties of metals

- ***Reactivity series*** is the arrangement of metals according to the reducing power(reactivity).
- The series was arranged by using the reaction of metals with water(H₂O), air (O₂) and dilute hydrochloric acid(HCl).

Metal	Reaction with air (O ₂)	Reaction with H ₂ O	Reaction with dil.HCl
K	Burn in air to form metal oxide. $4K + O_2 \rightarrow 2K_2O$	React with cold water to form hydrogen gas. $Na + H_2O \rightarrow NaOH + H_2$ $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$	React violently with dil. HCl $2Na + 2HCl \rightarrow 2NaCl + H_2$
Na			$Ca + 2HCl \rightarrow CaCl_2 + H_2$
Ca			
Mg		React with steam to form hydrogen gas. $Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$ $2Al + 6H_2O \rightarrow 2Al(OH)_3 + 3H_2$	React slowly down the list. $Zn + 2HCl \rightarrow ZnCl_2 + H_2$ $Fe + 2HCl \rightarrow FeCl_2 + H_2$
Al			
Zn			
Fe			
Pb	Form oxide when heated in air, but they don't burn. $2Pb + O_2 \rightarrow 2PbO$	They don't react with water at all.	
Cu			
Hg			
Ag			
Au	Not affected by air.		
Pt			

How does reactivity series affect extraction?

Increasing reactivity ↑

potassium
sodium
calcium
magnesium
aluminium
(carbon)
zinc
iron
lead
(hydrogen)
Copper*
mercury
silver
gold
platinum

Metals above carbon in the reactivity series must be extracted using **electrolysis (electrolytic reduction)**.

Metals less reactive than carbon can be extracted from their ores by **chemical reduction** using carbon, coke or charcoal.

Least reactive metals like platinum, gold, silver and copper can occur **native** (as free element) and they are just purified. But copper can be found combined state and extracted by chemical reduction.

Displacement reactions

- Takes place when most reactive element or radical takes place of least reactive one in a compound.
- Therefore the metals found at the top of the series tend to displace those found at the lower position in the reactivity series from the compound.
 - $2\text{Al} + \text{Fe}_2\text{O}_3 \rightarrow \text{Al}_2\text{O}_3 + 2\text{Fe}$
 - $\text{Zn} + \text{PbCl}_2 \rightarrow \text{ZnCl}_2 + \text{Pb}$
 - $\text{MgF}_2 + \text{Ag} \rightarrow \text{No reaction}$
- Ag is found below Mg in the reactivity series so it cant displace Mg from the compound, hence the reaction is not feasible.

Qn:

When iron metal is added to blue solution of copper (ii) sulphate, the solution changes to pale green, briefly explain why?

Ans:

◦ The solution changed to pale green due to formation of iron (ii) sulphate solution.



Blue **Pale green**

◦ Iron replaces of copper.

Try this:

◦ With the aid of a balanced chemical equation explain briefly why the blue solution of copper(ii) nitrate is decolourised when magnesium is added to it.

Extraction of metals

- Extraction of metal is the process whereby a pure metal is obtained from an ore.
- Ore is the naturally occurring solid from which metals are extracted. They are dug from the earth's crust and processed to obtain the metal.
- **Native** is the natural occurrence of a metal as an element (uncombined) in the environment.
- The branch of science concerned with the properties, production and purification of metal is called ***Metallurgy***.

Appearance of Different metals



Sodium



Potassium



Aluminum



Iron



Copper



Zinc



Methods of extracting metal

- There are two methods used in extracting metals
 1. ***Electrolytic reduction:*** This is applied for extracting the most reactive metals. Example: Sodium and aluminium.
 2. ***Chemical reduction:*** This is applied for extracting the moderately reactive metals. Example: Copper and zinc.
- The least reactive metals occur free, so they are just purified.
Example: Silver and gold.

Stages of extraction of metals

- There are four main stages of extraction of metals, namely
 1. Mining and concentration(dressing)
 2. Roasting in air
 3. Reduction of oxide to metal
 4. Purification of a raw metal

1.Mining and concentration(dressing)

- After obtaining the ore from the earth's crust, it is crushed into fine powder.
- The ore is fed into water tank containing frothing chemicals to remove the impurities. This is called froth flotation.
- Finally the ore is gently heated below it's melting point in absence of air to remove water and carbon dioxide from carbonates.

2. Roasting in air

- The heated ore in presence of air to get metal oxide.
- This is mostly done for the ores which occur as sulphide.

3. Reduction of oxide to metal

- The metals which are moderately reactive are reduced from their oxides to form a metal.
- This can be done by using carbon (coke) or carbon monoxide.

4. Purification of a raw metal

- This is the last stage where pure metal is to be obtained.
- Purification can be done in several ways depending on the properties of the metal and the impurities in it.
- The following are the methods of purification
 - i) Distillation
 - ii) Oxidation
 - iii) Electrolysis

i) Distillation

- The crude metal is heated in the furnace until the metal evaporates leaving the impurities behind.
- The vapour of the metal is condensed and collected in separate chamber.
- This can easily be done if the metal is more volatile than the impurities. Example: Mercury and zinc

ii) Oxidation

- Molten crude metal is exposed in air in a furnace.
- The impurities are oxidized and escape as gas or form scum (a layer of dirtiness or froth) over molten metal and removed by skimming(taking the floating layer).
- This method is only used when the impurities have greater affinity to oxygen than the metal.

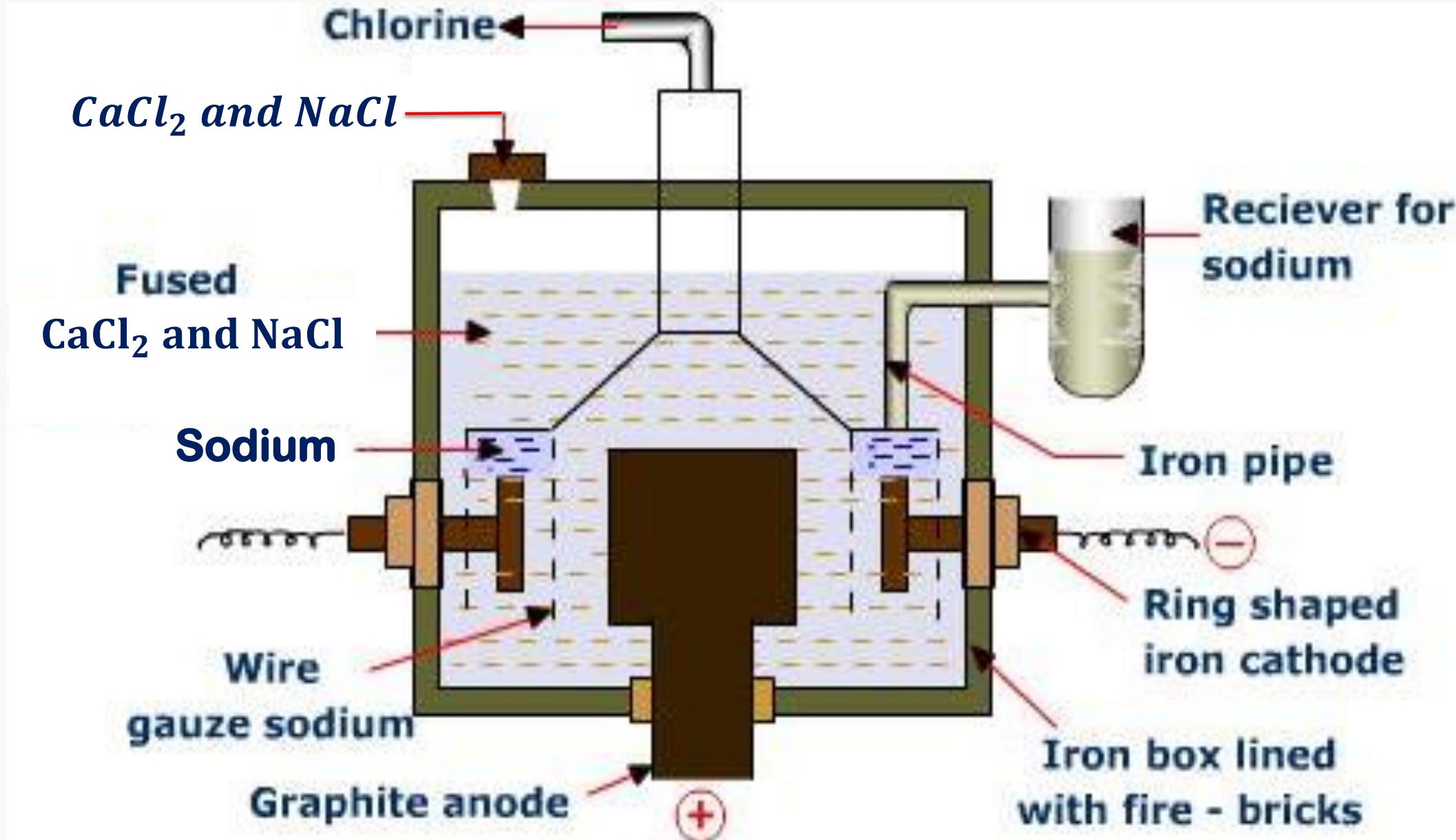
iii) Electrolysis

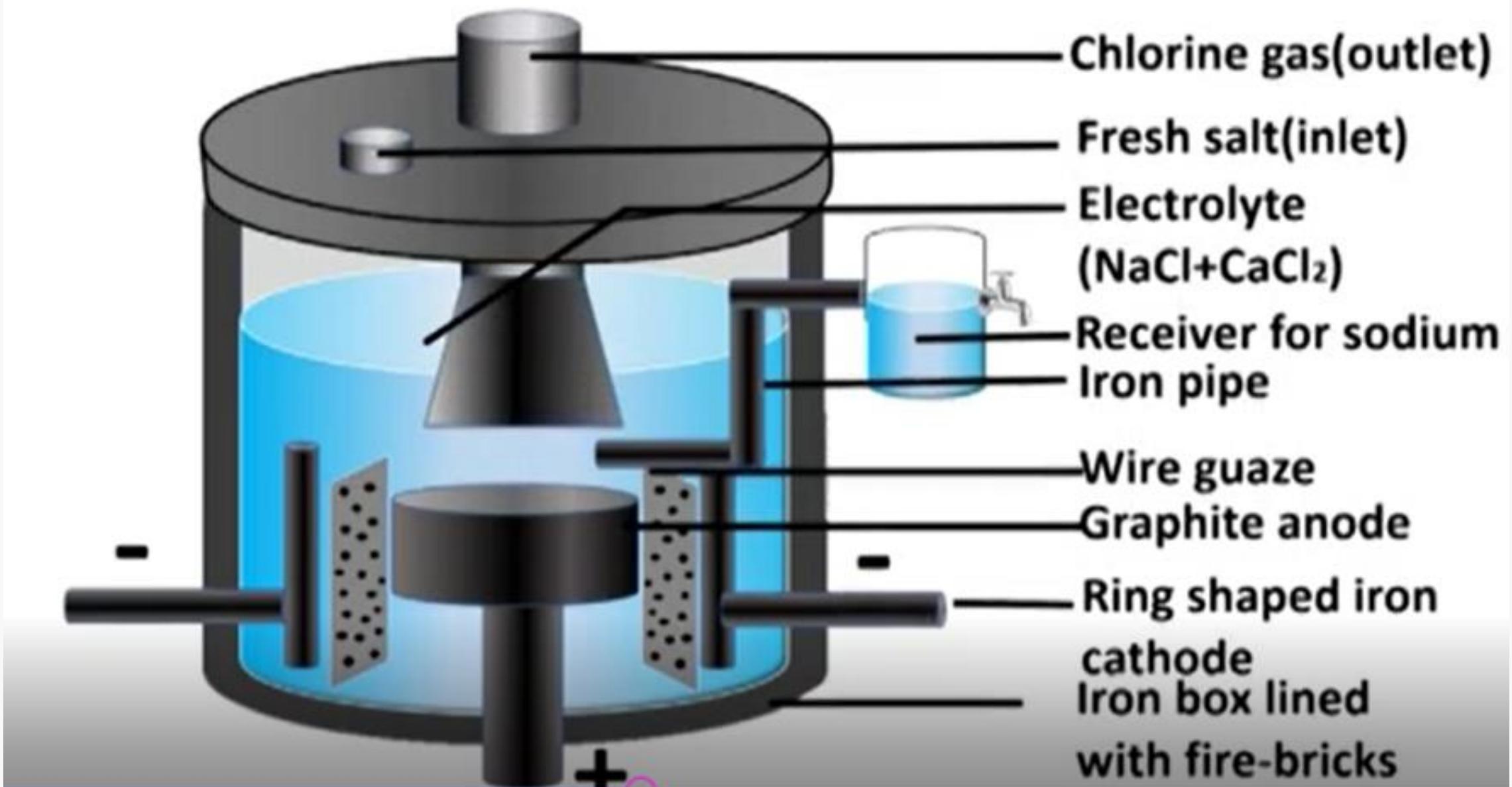
- The crude metal is molded into a block and placed as an anode in electrolytic cell.
- The cathode is usually made up of a thin plate of the pure metal.
- The electrolyte used should be the one which contains the ions of the particular metal. *Example:* Copper and silver
- For example in purification of copper
 - **Anode:** Crude copper
 - **Cathode:** Pure copper
 - **Electrolyte:** Copper(ii) sulphate

1) EXTRACTION OF SODIUM

- Sodium does not occur naturally as a free element. This is because it is too reactive.
- The chief ore of sodium is ***common salt*** (NaCl) which occurs naturally in the earth's crust.
- Since sodium chloride does not contain sulphur it is not roasted.
- Sodium is extracted by electrolytic reduction because it is one of most reactive metals.
- This electrolytic reduction to obtain sodium is called **Down's process**, and the electrolytic cell that is used in this process is called **Down's cell**.

The down's cell





Down's process

- Sodium chloride is electrolyzed when it is in molten condition.
- The ore is crushed and mixed with calcium chloride in the ratio of $NaCl: CaCl_2 = 2: 3$ and put into the Down's cell.
- The function of calcium chloride in the Down's cell is to lower the melting point of the sodium ore from $800^{\circ}C$ to $600^{\circ}C$.
- During this process chlorine is produced at the anode.
- Sodium rises up due to its density and tapped off through the iron vessel placed over the cathode.
- The wire gauze (diaphragm) between anode and cathode prevents recombination of sodium and chlorine when they escape.

- Sodium which has collected contains calcium which is also produced during electrolysis, but calcium crystallizes after cooling and pure sodium is obtained.

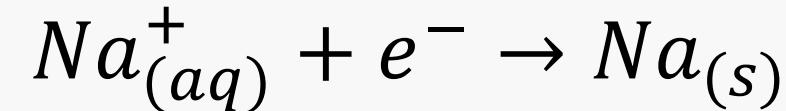
At anode:

- Here chloride is discharged losing two electrons and chlorine gas is produced and escape through the hood.



At cathode:

- Here sodium is discharged by gaining one electron and collects in the inverted iron trough.



Uses of sodium

- i) Used in making alloys.
- ii) Used as table salt.
- iii) Used for making baking powder.
- iv) Used to make sodium fluoride which makes toothpaste.
- v) Used in different health therapies. Example: Rehydration therapy $NaCl$, KCl and $C_6H_{12}O_6$ (glucose).
- vi) Used to make sodamide and cyanide which are used to isolate gold from metallic ores.

NOTE: Sodium and most of other group I elements are kept in paraffin oil because the oil is less dense than sodium thus it the metal from being contaminated.

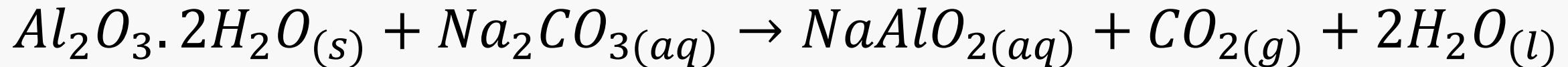
2) EXTRACTION OF ALUMINIUM

- Aluminium is the most abundant metals on the earth's crust (about 8% of earth's crust).
- Aluminium always occur state, the following are the ores in which it occurs:
 - ❖ Bauxite (Al_2O_3)
 - ❖ Cryolite (Na_3AlF_6)
 - ❖ Felspar ($KAlSi_3O_8$)
 - ❖ Kaolin ($Al_2Si_2O_7 \cdot 2H_2O$)

- Among the ores above, the bauxite and cryolite are the chief ores from which aluminium is extracted although bauxite is mostly used.
- Being one of the more reactive metals, aluminium is extracted by electrolysis.
- The method of extraction of aluminium is called Hall-heroult method.
- The process of extraction involves
 - i. Purification of the bauxite
 - ii. Electrolysis of alumina

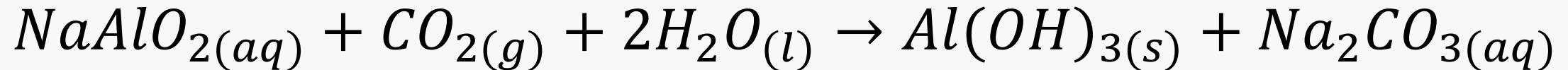
Purification of the bauxite

- This is the process where the impurities Fe_2O_3 and SiO_3 impurities through Hall process.
- The bauxite is powdered and fused with sodium carbonate where by the bauxite dissolves and form sodium aluminate.
- In this process the impurities Fe_2O_3 and SiO_3 are left unaffected.



- The molten mass is removed with water leaving the impurities behind.

- Sodium aluminate is then heated to 50°C or 60°C in presence of carbon dioxide(CO_2) to form aluminium hydroxide precipitates.



- The product obtained is filtered to obtain aluminium hydroxide then washed and dried.
- The hydroxide obtained is calcinated (heated strongly) at 1500°C to form alumina.



- The Alumina produced is ready to be put into the electrolytic cell for electrolysis.

- Aluminium can not be extracted by reduction of alumina using carbon(coke), because it has higher affinity to oxygen than carbon.
- However at high temperature aluminium reacts with carbon to form aluminium carbide(Al_4C_3).
- Therefore electrolysis is the suitable method to obtain aluminium.

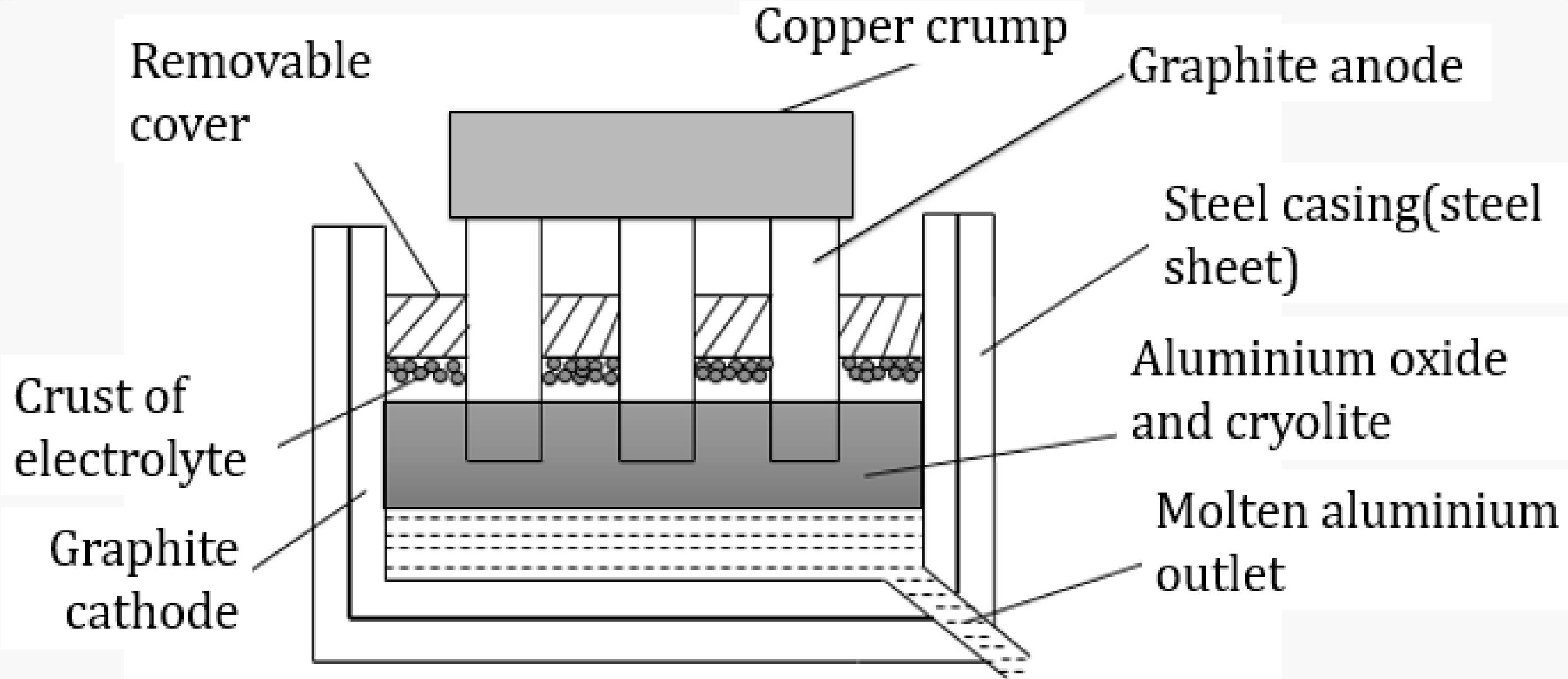
Electrolysis of alumina(Al_2O_3)

- Alumina is mixed with cryolite and are put together into the electrolytic cell. A lot of electricity is required in this process.
- Once aluminium is produced, it is collected at the bottom of the cell in the molten state due to it's high density.
- The temperature of the cell is maintained at $800^{\circ}C - 900^{\circ}C$. But the melting point of alumina is over $2000^{\circ}C$.

The functions of cryolite

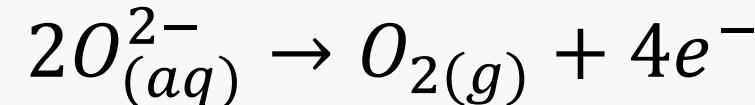
1. To lower the melting point of alumina.
2. To improve electrical conductivity of alumina (It is poor conductor by nature).

Extraction of aluminium by electrolysis.mp4



At anode:

- Here oxide is discharged losing two electrons and oxygen gas is produced.



- The oxygen gas produced reacts with carbon anodes forming carbon dioxide gas, this cause the anode to burn away and hence need to be replaced regularly.



At cathode:

- Here aluminium is discharged by gaining three electrons and collected at the bottom.



Uses of aluminium

- i. Used in manufacturing of alloys eg: *bronze* (Al&Cu), *magnalm* (Mg&Al).
- ii. Used as reducing agent.
- iii. Used to manufacture house hold utensils.
- iv. Used in making mirrors.
- v. Used in making heat resistant clothing for fire fighting, due to it's reflectivity.
- vi. Used in making overhead power cables, due to it's good electrical conductivity.

vii. Used in making roofing materials, due to it's corrosion resistance.

Qn:

◦ Extraction of alumunium is said to be very expensive, why?

Answer :

- A lot of electricity is used to extract it.
- Oxygen released at the anode reacts with carbon and cause it to burn away.

Qn: Give reason to why aluminium is chosen to make each of the following items;

- i. Cooking foil
- ii. Overhead power cables
- iii. Windows frames

Answer:

i. Aluminum is used to make cooking foil because:

- It provides a ***complete barrier*** to light, oxygen, moisture and bacteria. For this reason, foil is used extensively in ***food*** and ***pharmaceutical*** packaging.
- It is a good conductor but a very poor radiator and absorber because it is shiny.

ii. Aluminum is used to make Overhead power cables because:

- It is cheap
- It is light hence do not need large support structures.

iii. The advantage of choosing **aluminium windows** and doors is due to

- The strength of **aluminium**,
- Less material is required in the **frame** to hold the glass
- Its shiny in nature.
- It is not affected by UV light rays.

3) EXTRACTION OF IRON

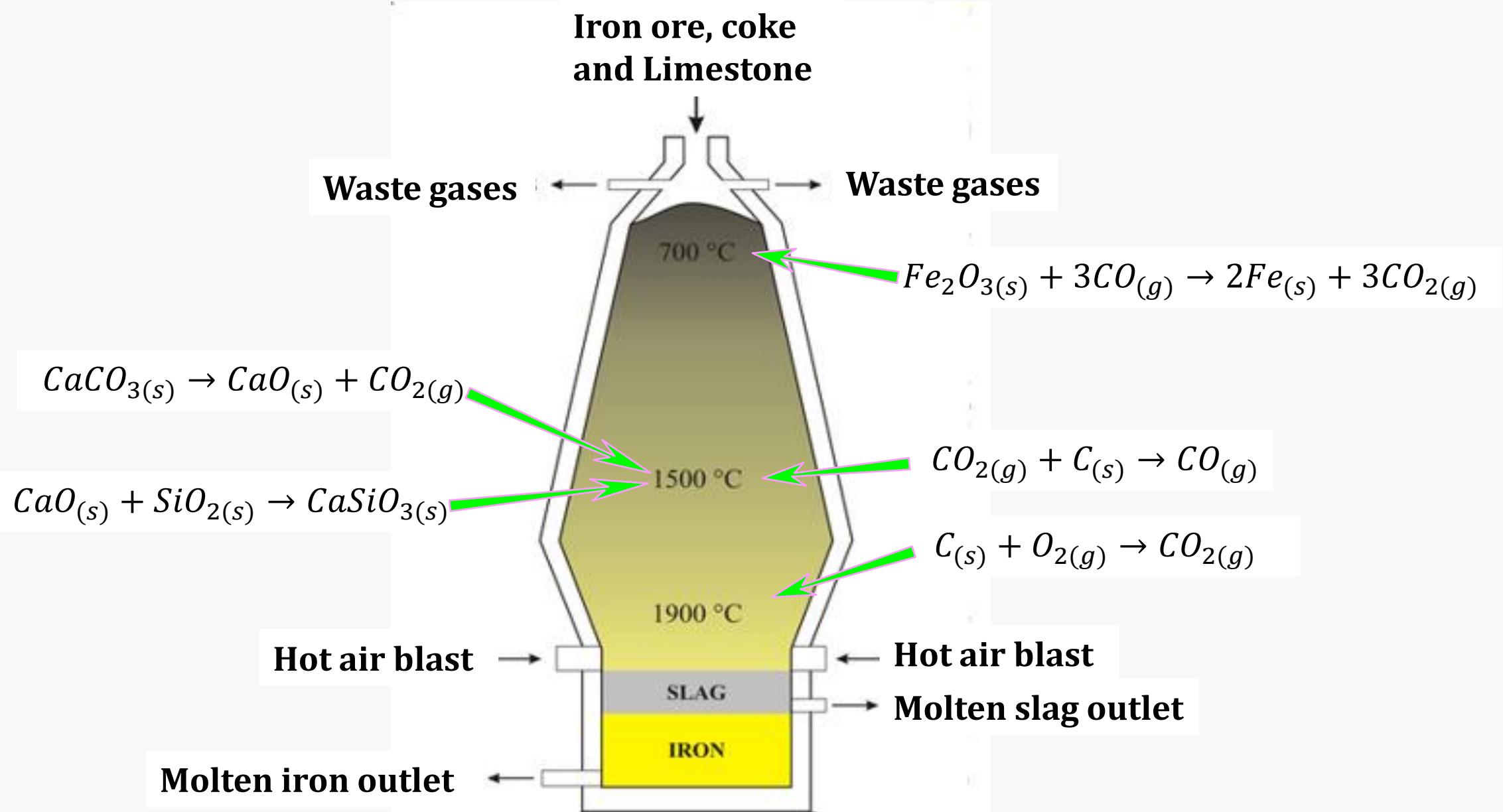
- Iron is the second abundant metal occurring on the earth's crust after aluminium.
- Iron is relatively reactive metal, so it occurs in combined state.
- The iron ores are;
- Haematite (Iron(iii) oxide, Fe_2O_3)
- Magnetite (Triiron tetraoxide, Fe_3O_4)
- Spathic/siderite(Iron(ii) carbonate, $FeCO_3$)
- Iron pyrites (Iron disulphide, FeS_2)
- Limonite (hydrated iron(ii) oxide, $Fe_2O_3 \cdot H_2O$)

- The chief iron ores of iron are haematite and magnetite which contains about 70% iron and 72.4% irons respectively.
 - Iron pyrites are mostly used in manufacturing sulphuric acid rather than extraction of iron.
- ## ◦ **Concentration**
- After mining the ore is dressed to remove earth's impurities such as sand, clay and other non-magnetic impurities.
 - This involves crushing and washing the ore.

In the blast furnace

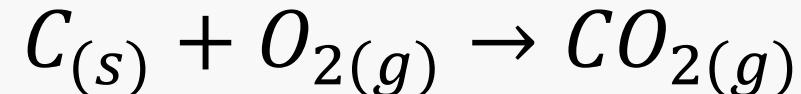
- **Blast furnace** is a tall oven used to extract iron from iron ore by burning it with carbon at high temperatures.
- **Iron ore, limestone** and **coke** are all put from the top of the blast furnace.
- The function of limestone is to remove the silica impurities(SiO_2)
- The function of coke is to reduce the oxide of iron.
- The blast furnace has pipes at the base for passing **hot air blast** into the furnace.
- At its top part has the pipes which pass out **waste gases**.

The Blast Furnace

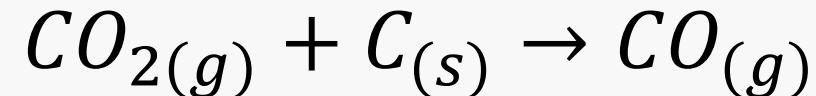


Reactions in the blast furnace

- When hot air enters it oxidizes the coke at the lower part of the furnace to form carbon dioxide.



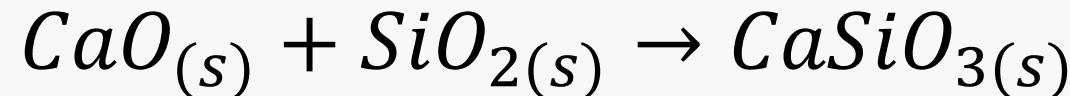
- This reaction liberates large quantity of heat which keeps high temperature necessary for reduction process.
- Then carbon dioxide rise up and react with more coke in limited supply of air to form carbon monoxide.



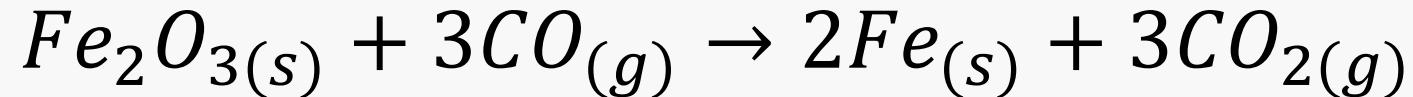
- At the same temperature, calcium carbonate (lime stone) decomposes to form calcium oxide.



- Then calcium oxide produced reacts with silica impurities to form calcium silicate (slag) in molten form.



- The carbon monoxide formed above rises up and reduce iron(iii) oxide to form iron and carbon dioxide.



- Carbon dioxide escapes as waste gas at the top of the furnace.
- The temperature 300°C is too low to melt iron, the metal obtained is called spongy iron.
- Spongy iron sinks at the bottom hotter part and melt at about 1300°C and the molten iron collects at the hearth(floor of the fire place)

- Molten slag is less dense than iron hence it floats over molten iron forming separate layer.
- A layer of slag protects iron from being oxidized by the hot air.
- The mixture of waste gases (N_2 , CO_2 and CO) are collected for heating air blast.
- The iron obtained is impure, it is called *pig-iron* or *cast iron* (but mostly pig-iron).

Uses of iron

- Used to make alloy
- Used to make water pipes.
- Used to make building materials.eg nails, iron rods and sheets.
- Used for making simple and complex machines.eg drillers, loader
- It is a component of chlorophyll, haemoglobin and myoglobin.
- Used in rail way construction.

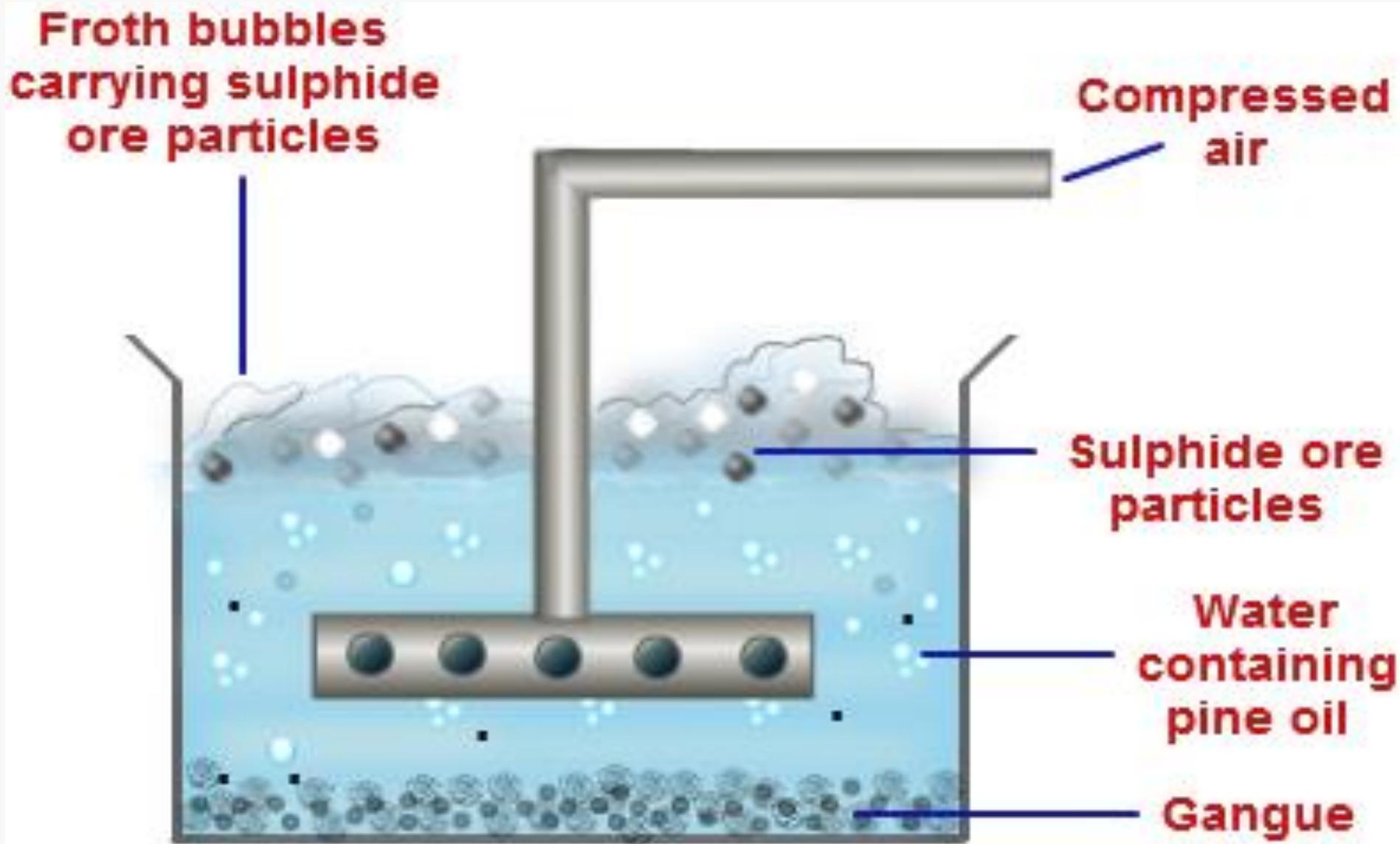
4) EXTRACTION OF COPPER

- Copper is the second oldest metal after gold.
- It was given the name cuprum because it was discovered in Cyprus island by that time.
- Copper occurs in both free and combined state.
- Main ores of copper are
 - Copper pyrites ($CuFeS_2$)
 - Copper glance/Copper(i) sulphide (Cu_2S)
 - Cuprite/Copper(i) oxide (Cu_2O)
 - Malachite ($CuCO_3 \cdot Cu(OH)_2$)
- Copper is mainly extracted from copper pyrites by dry process.

Concentration

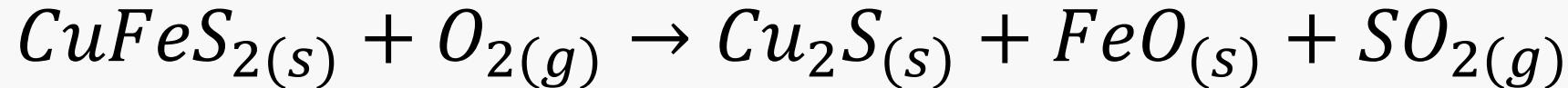
- After mining the ore is finely divided and concentrated by froth flotation.
- In this process the crushed ore is suspended water containing pine oil (frothing chemicals).
- A blast of air is then passed through the suspension
- The particles of the ore get wetted by oil and float as froth, which is then skimmed to separate it from impurities.
- The gangues (impurities) sink to the bottom.

Froth flotation



Roasting in air

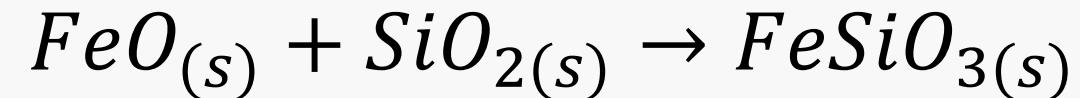
- Concentrated ore is roasted in the furnace in presence of excess air.
- During this process the impurities such as Sulphur, arsenic and antimony are oxidized and removed as gases.



Impurities:

- $S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$
- $As_{(s)} + O_{2(g)} \rightarrow As_2O_{3(g)}$
- $Sb_{(s)} + O_{2(g)} \rightarrow Sb_2O_{3(g)}$

- Then by adding silicon dioxide (SiO_2) and heating in absence of air, the iron(ii) oxide is converted to slag of iron(ii) silicate($FeSiO_3$) and separated from copper(i) sulphide (Cu_2S).



Reduction

- The copper(i) sulphide is then reduced to copper metal by heating in a regulated supply of air.



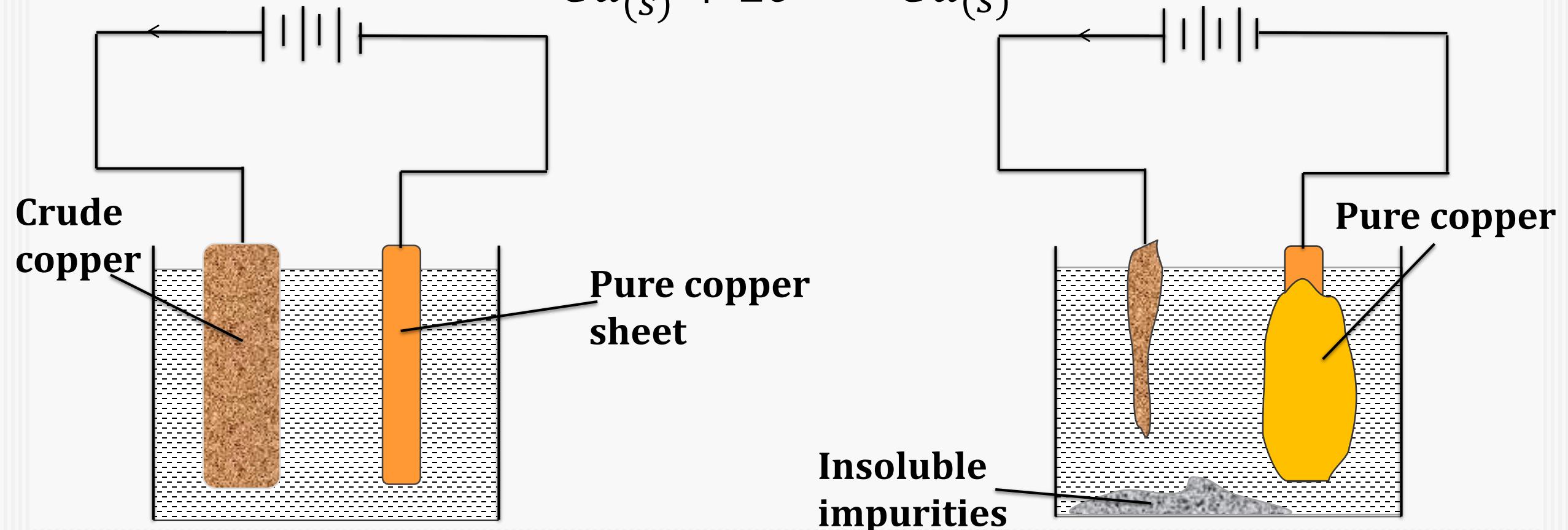
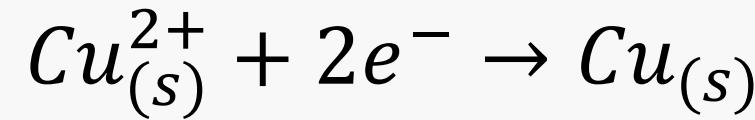
- This copper obtained is too impure for many of the purposes to be used in industry. Therefore it should be refined.

Purification of copper

- The best way of purifying copper is through electrolysis.
 - Anode: Impure copper
 - Cathode: Pure sheet of copper
 - Electrolyte: Copper(ii) sulphate (few drops of H_2SO_4 are added to increase conductivity).
- During electrolysis the impure copper anode dissolve losing two electrons and form copper ions in the electrolyte.



- Then pure copper deposits at the cathode. The copper deposited at the cathode is pure.
- but the soluble impurities dissolve in the electrolyte and insoluble will at the bottom. Purifying copper.mp4



Uses of copper

1. Used in alloying eg: brass (Cu&Zn), bronze (Cu, Sn, Al, Mn&Zn).
2. Makes water pipes and boiler.
3. Used as lightening rod.
4. Used in roofing.
5. Used as wood preservative reagent when liquefied.
6. Used in construction of machines like trains and cars.
7. Used in making calorimeters
8. It serves as enzyme co-factor in living organisms.
9. It avoid hair problems.
- 10.Used in making electrical wires and cables.
- 11.It improves elastic fibres in skin.

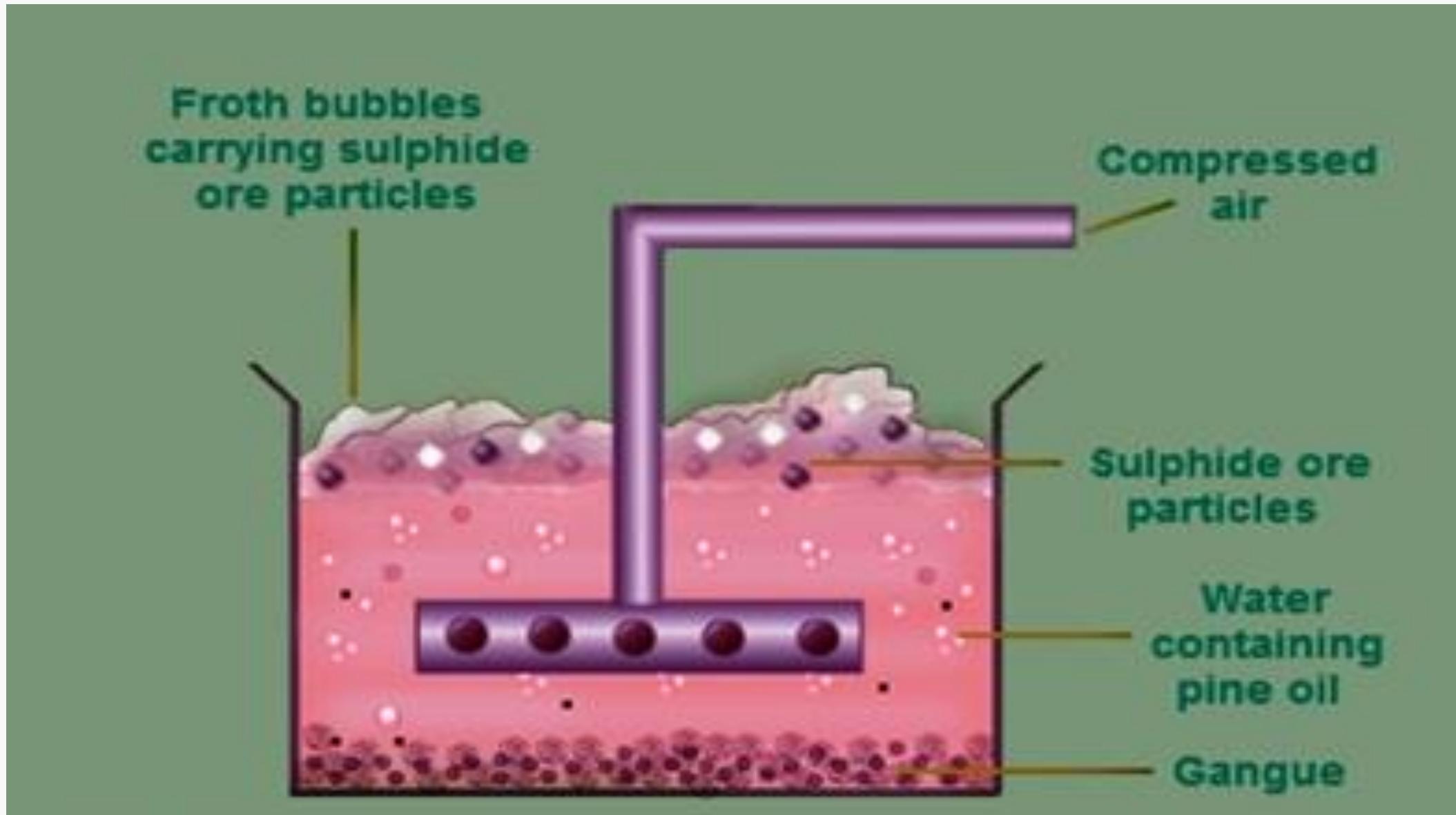
5) EXTRACTION OF ZINC

- Zinc is one of the moderately reactive metals.
- It is extracted by chemical reduction method by the instrument known as fire-clay retort.
- The ores of zinc are
- Zinc blonde/Zinc sulphide (ZnS)
- Calamine/Zinc carbonate ($ZnCO_3$)

Concentration

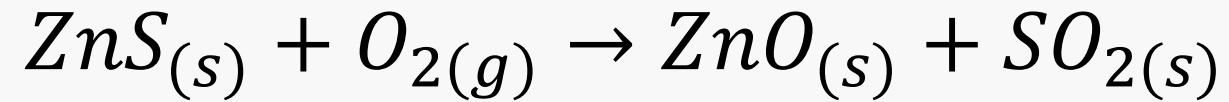
- After mining the ore is finely divided and concentrated by froth flotation.
- the crushed ore is suspended water containing frothing chemicals/collectors(pine oil or fatty acids).
- The particles of the ore get wetted by oil and gangues wetted by water. Then the froth of the ore is removed by skimming.

Froth flotation of zinc blende



Roasting in air

- A concentrated ore is then suspended in a rising stream of air and zinc blende is converted to zinc oxide.



Reduction

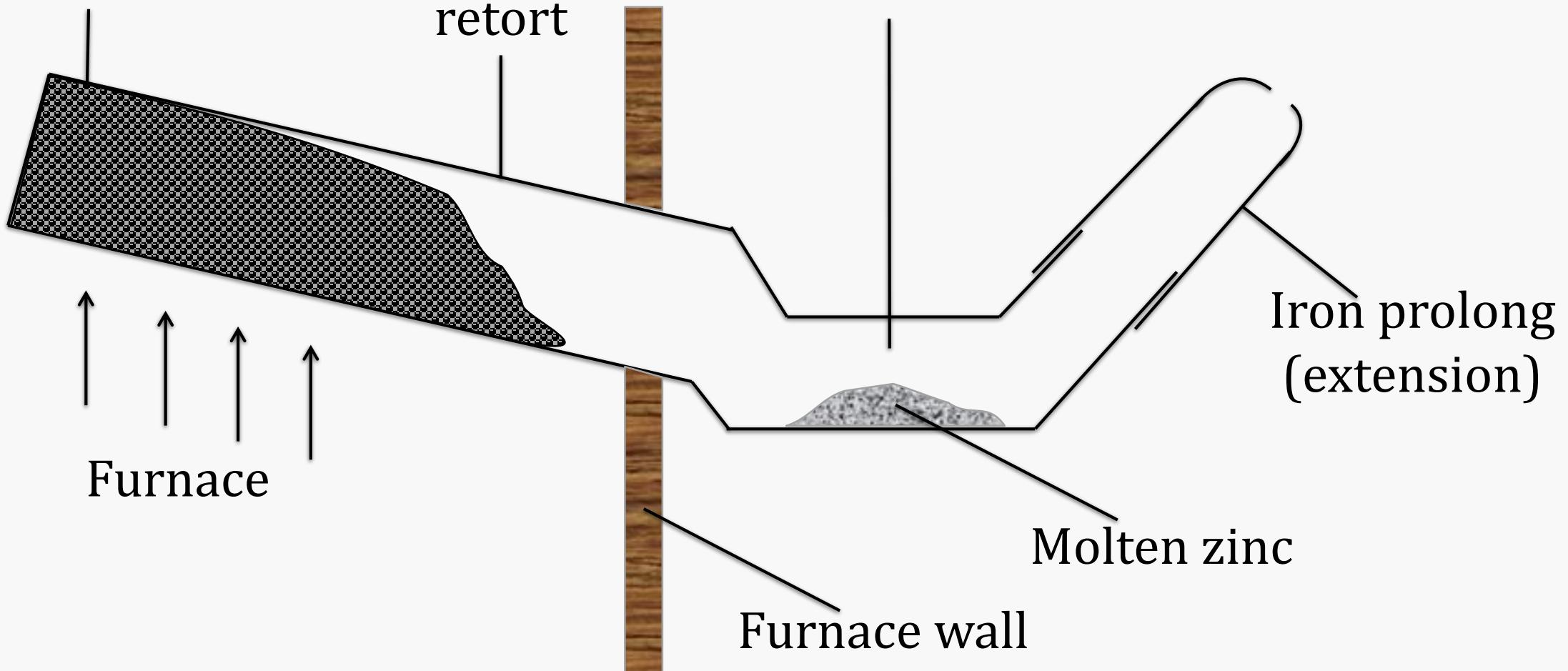
- Zinc oxide is mixed with coke and put together in the fire-clay retort and heated at 1400°C.
- This process is called Belgian process.
- Here zinc oxide is reduced by coke to zinc.



- At the end of the fireclay retort there is fireclay condenser which is connected with iron prolong.
- The iron prolong collects any zinc which escapes the condenser.

The fireclay retort

Mixture of zinc oxide and coke



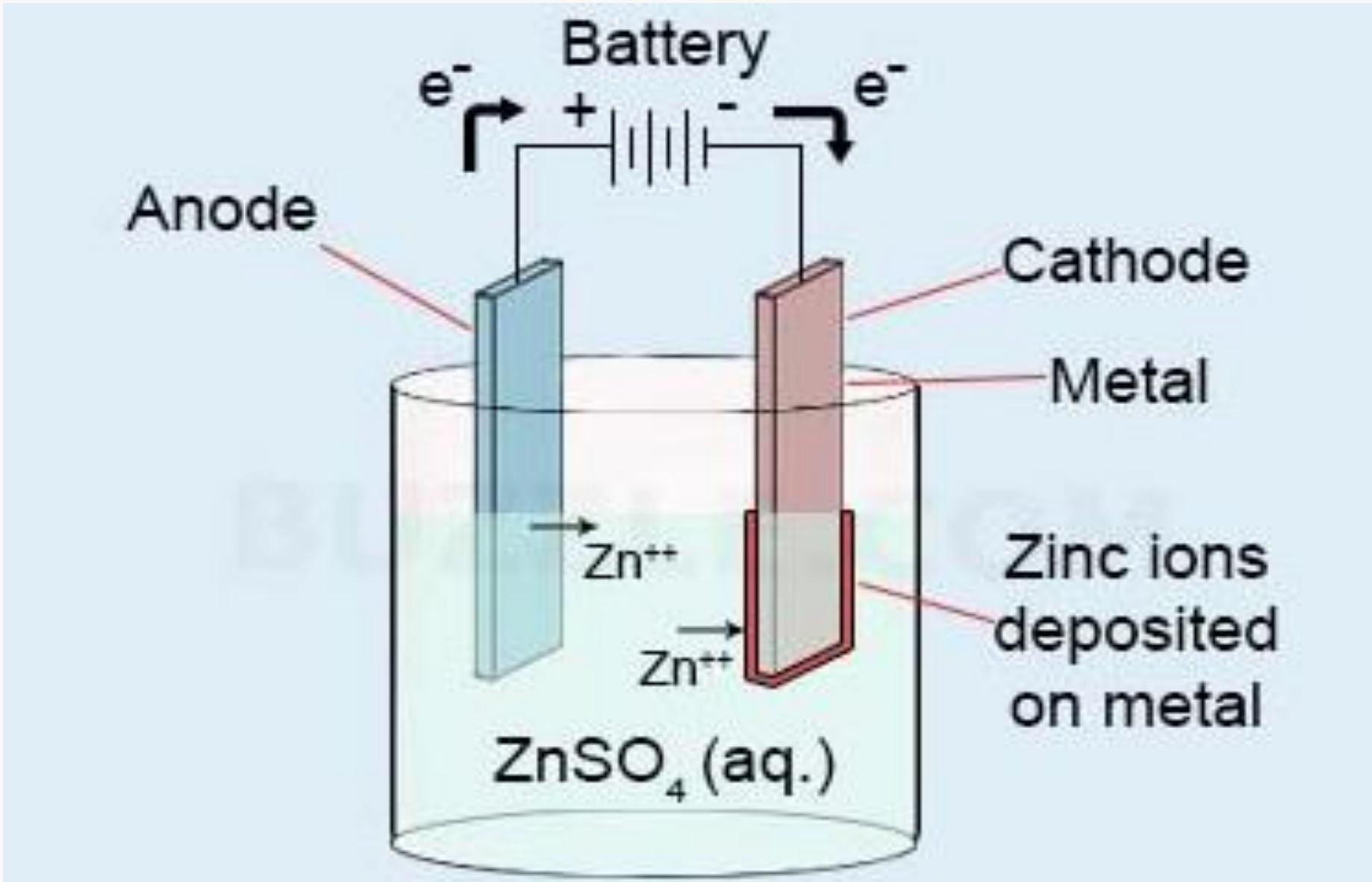
Purification of zinc

- Zinc obtained from the fireclay retort is impure, so it must be purified.
- Due to its low boiling point(907°C), zinc can be purified by fractional distillation.

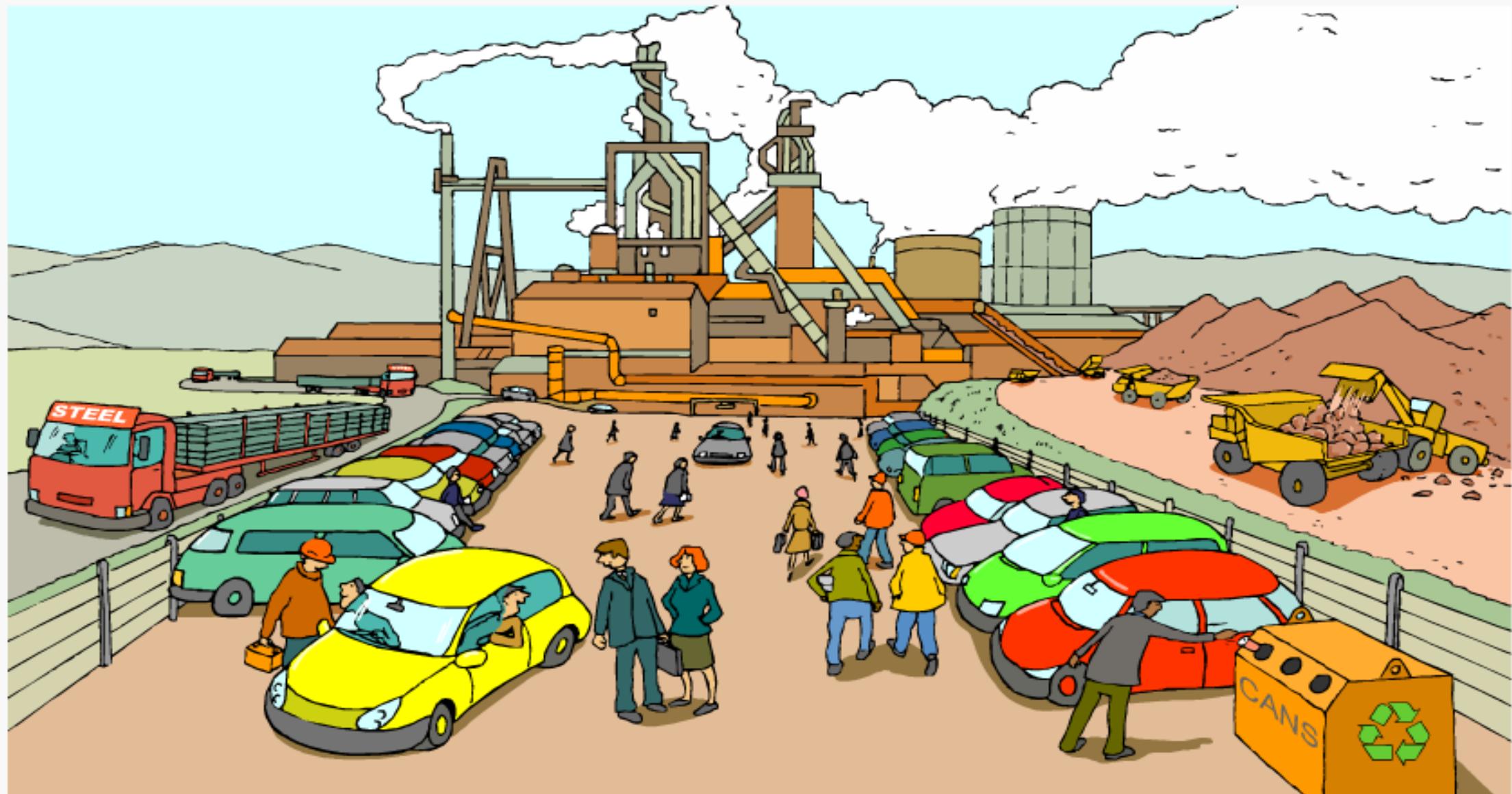
Uses of zinc

- Making alloys example: brass and bronze
- Used in galvanization.
- Used in dry cells.

Purification of zinc



Metal extraction is a large and important industry.
What positive and negative impacts does it have?



Environmental problems associated with mining and extraction of metals

1. Mercury and lead compounds cause nerve deformity in unborn baby and serious nerve disorder in adults.
2. Production of acidic gases (SO_2 and CO_2) causes acidic rain.
3. Explosive remains may destroy habitat of living organisms.
4. Rock fragments and soil heaps spoil the beautiful natural appearance of landscape.
5. Aluminium leads to early onset of brain diseases(Azheimer's brain disease)

6. Cavities left after mining lead to water logging and flooding.
7. Cyanide used in extraction of metals when poorly disposed may affect both aquatic and terrestrial life.
8. Due to poor methods used in local mines, many accidents occurs.

Measures to reduce problems associated with mining and extraction of metals

1. Planting trees to replace those destroyed during mining.
2. Installing scrubbers in industries to absorb pollutant gases.
3. Mines should not be established near by water sources.
4. Heaps should be leveled and holes should be rehabilitated to improve physical features of the area.
5. Chemicals used in extraction should be detoxified before disposing them.
6. Local mining should be done with precautions to avoid unnecessary accidents.

TOPIC 9

COMPOUND OF

METALS

Compound

- A compound is a substance formed by the chemical Combination of two or more elements.
- Metals react with other substances, especially non- metals, to form different compounds. These are called compound of metals.

◦ The compounds of metals include:

- i) Metal oxides
- ii) Metal hydroxides
- iii) Metal carbonates
- iv) Metal Hydrogen Carbonates
- v) Metal nitrates
- vi) Metal chlorides
- vii) Metal sulphates

9.1 METAL OXIDE

- Metal oxide is the binary compound which contains a metal and oxygen.

Preparation of metal oxides

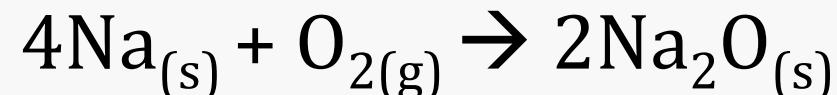
- Metal oxides can be prepared using two methods namely:
 - i) Direct method
 - ii) Indirect method

a) Direct Method

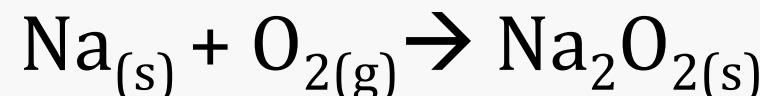
- In the direct method, a metal is directly heated in oxygen to give a metal oxide.

Example:

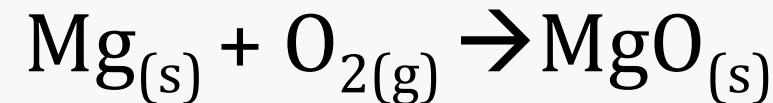
- ❖ When sodium is burned in limited supply of air (or just exposed to air) it reacts with oxygen to form sodium oxide which is a ***whitish-grey solid***. It burns with ***golden yellow flame***.



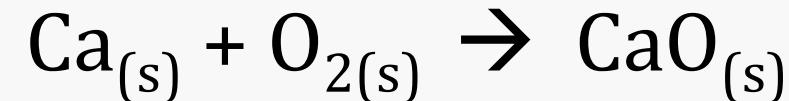
- ❖ In plenty supply of air it forms sodium peroxide which is a ***pale yellow solid***.



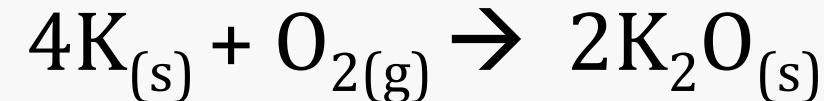
- Magnesium burns in air with a *very brilliant white flame* to form a **white solid** of magnesium oxide



- Calcium burns in air with a *brick-red flame* to form calcium oxide which is **white solid**



- Potassium burns in limited supply of air with *lilac flame* to form **white solids** of potassium oxide.



b) Indirect Method

- This involves decomposition of carbonates, hydroxides and nitrates of some metals.



NOTE: Oxides of **Al** and **Be** can only be prepared by indirect method. This is because when they are heated they forms a **hard protective oxide coat** which prevents further reaction.

Classification Of Metal Oxides

- Metal oxides can be categorized based on the following
 - a) Nature of an oxide
 - b) Solubility of an oxide in water

A) Nature Of An Oxide

- According to nature of an oxide we have ***basic oxides*** and ***amphoteric*** oxides

i) ***Basic Oxides***

- These are the oxides which react with acids to give salt and water only. They also form hydroxide when dissolved in water.
- Basic metal oxides include all the oxides of Group I metals and some of group II metals like Ca and Mg.

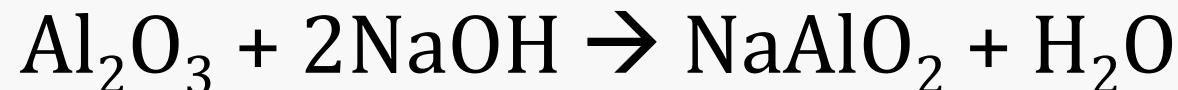


ii) Amphoteric Oxide

- Are the oxides which shows both basic and acidic properties.
 - They reacts with both acids and bases and form salts.
 - Example: Aluminium oxide (Al_2O_3), Berylium oxide(BeO), zinc oxide(ZnO), lead(ii) oxide(PbO), tin oxide(SnO_2) and copper(ii) oxide(CuO).
- ❖ With acids they form salt and water.



- ❖ With base they form complex salt and water.



Other Types Of Oxides

a) Acidic oxide

- These are the oxides of non-metals which when dissolved in water they form acidic solution.

Example: CO₂, SO₂, SO₃, and NO₂



b) Neutral oxides

- Are the oxides which have neither basic nor acidic properties.
- They don't react with acid or base.

Example: CO and NO

c) Mixed oxides

- Are the oxides of two simple oxides.

Example: Fe_3O_4 Mixture of FeO and Fe_2O_3

N_2O_5 Mixture of NO and N_2O_3

d) Peroxides

- Are the oxides formed when an element is burned in excess air.

Example: Na_2O_2 , H_2O_2 , N_2O_2

B) Solubility Of An Oxide in Water

- According to solubility of an oxide in water there are two types of oxides which are ***soluble oxides*** and ***insoluble oxides***.

i) ***Soluble oxides***

- These are the oxides which dissolves in water.

Example: Oxides of group I elements. Oxide of Ca is slightly soluble.



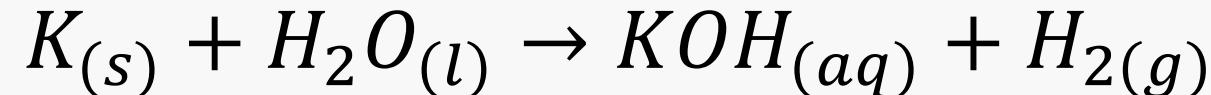
ii) Insoluble oxides

- Are the oxides which do not dissolve in water.
- Most metal oxides are insoluble in water.

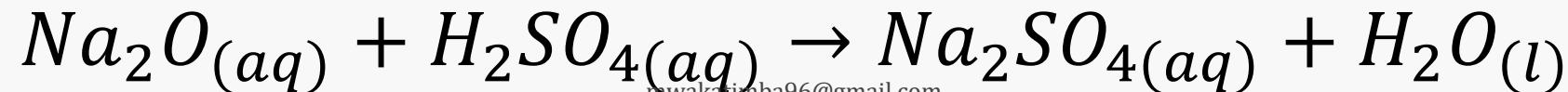
Example: Aluminium oxide, zinc oxide, lead (II) oxide, copper (II) oxide and iron (II) and (III) oxides do not dissolve in water.

Properties Of Metal Oxides

- **Solubility:** With exception of oxides of Na, K, Ba and Sr other metal oxides are insoluble in water.
- **Action of heat:** All metal oxides are stable on heating, they doesn't decompose.
- **Reaction with water:** Soluble metal oxides react with water to form metal hydroxides.



- **Reaction of with acids:** Metal oxides react with dilute acids to form salt and water only.



Uses of Metal oxides

- Calcium oxide is used as drying agent. Example in drying ammonia gas.
- calcium oxide used in liming.
- Calcium oxide is used to trap impurities in the blast furnace.
- Calcium oxide is used is used in manufacturing cement.

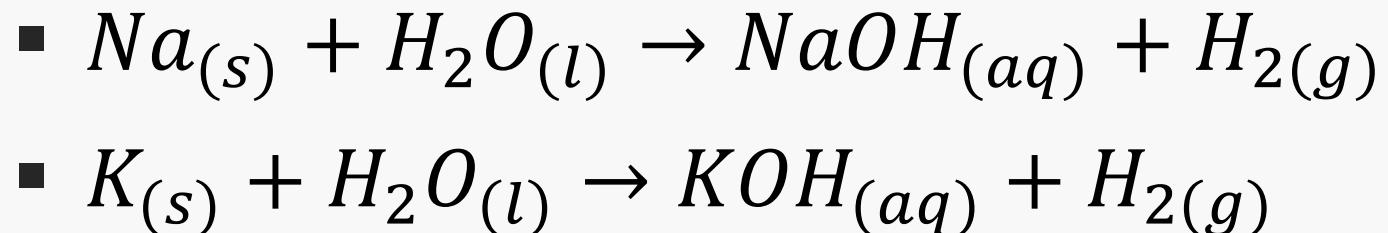
9.2 METAL HYDROXIDE

- Metal hydroxides is any inorganic compound that contains the hydroxyl group (-OH).
- The metal Hydroxides are bases.

Preparation of metal hydroxide

i) Direct Method

- The Metals those have high in the Reactivity Series are reacted with water to form the Hydroxide and hydrogen gas.



ii) Indirect method

- The Hydroxide is obtained by reaction between Alkalies (Base) with aqueous salts to form Hydroxide.
- This is mostly used to prepare insoluble metal hydroxides.
- In this reaction the soluble hydroxides are used as precipitating reagents
- $Pb(NO_3)_3(aq) + 2NaOH_{(aq)} \rightarrow Pb(OH)_{2(s)} + 2NaN_3(aq)$
- $Cu(NO_3)_2(aq) + 2NH_4OH_{(aq)} \rightarrow Cu(OH)_{2(s)} + 2NH_4NO_3(aq)$
- $Al(NO_3)_3(aq) + 3NaOH_{(aq)} \rightarrow Al(OH)_{3(s)} + 3NaN_3(aq)$

Classification of Metal Hydroxides

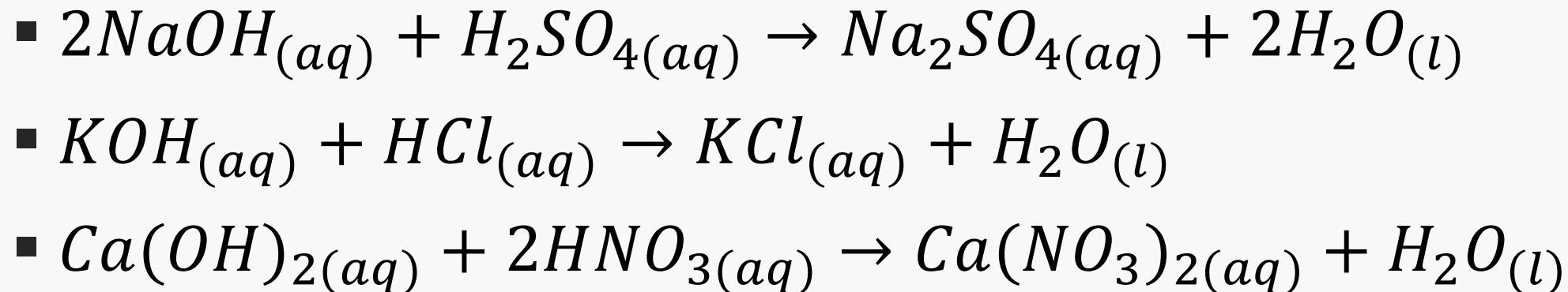
- Metal hydroxides can be categorized based on the following
 - a) Nature of hydroxide
 - b) Solubility of an oxide in water

A) Nature Of hydroxide

- According to nature of hydroxide we have ***basic hydroxides*** and ***amphoteric hydroxides***.

i) ***Basic hydroxides***

- Basic hydroxides are those which react with acid to form salt and water. They show basic characteristics only.



ii) Amphoteric hydroxides

- Amphoteric hydroxides are those which can react with both acid and base. They show both acidic and basic nature.
- Example: Hydroxides of aluminium, beryllium zinc and Lead.

With acid

- When react with acid they form salt and water.

- $Zn(OH)_{2(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$
- $Al(OH)_{3(s)} HCl_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(l)}$
- $Be(OH)_{2(s)} 2HNO_{3(aq)} \rightarrow Ca(NO_3)_{2(aq)} + H_2O_{(l)}$

With base

- When excess Hydroxide of Sodium or potassium is added to amphoteric metal hydroxides, they dissolve to form Complex ions.

- $Zn(OH)_{2(s)} + NaOH_{(aq)} \rightarrow Na_2ZnO_{2(aq)} + H_2O_{(l)}$
- $Al(OH)_{3(s)} + KOH_{(aq)} \rightarrow KAlO_{2(aq)} + H_2O_{(l)}$
- $Be(OH)_{2(s)} + 2NaOH_{(aq)} \rightarrow Na_2BeO_{2(aq)} + H_2O_{(l)}$

B) Solubility Of hydroxide

- According to solubility of hydroxide there is ***soluble hydroxides*** and ***insoluble hydroxides***.

i) ***Soluble hydroxides***

- These are the metal hydroxides which dissolves in water.
- They don't react with water but the only dissociate.
- There are very few soluble hydroxides, these are;
- $NaOH$ and KOH , but $Ca(OH)_2$ is slightly soluble in water.

ii) Insoluble hydroxides

- These are the metal hydroxides which do not dissolve in water.
- They don't react with water but they only dissociate.
- There are many insoluble hydroxides.
- All metal hydroxides are insoluble in water except those mentioned above.
- **Example:** $Be(OH)_2$, $Al(OH)_3$, $Zn(OH)_2$, $Mg(OH)_2$, $Pb(OH)_2$.

Properties of Metal Hydroxides

1. **Solubility:** All metal hydroxides are insoluble in water except those of sodium, potassium, lithium and calcium.
2. **Action of Heat:** Hydroxides of Na & K do not decompose on heating but the rest decompose to give a metal oxide and water.
 - $KOH_{(s)} \rightarrow$ No reaction
 - $Al(OH)_3{}_{(s)} \rightarrow Al_2O_3{}_{(s)} + H_2O_{(l)}$
3. **Action on Acids:** Metal hydroxides react with mineral acids to give a Salt and Water.
 - $2NaOH_{(aq)} + H_2SO_4{}_{(aq)} \rightarrow Na_2SO_4{}_{(aq)} + 2H_2O_{(l)}$
 - $KOH_{(aq)} + HCl_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(l)}$

Uses of Metal Hydroxides

1. Calcium Hydroxide (slake lime) Uses as Liming Material in soil treatment.
2. Manufacture of paints
3. Uses in Qualitative Analysis
 - $NaOH$ is used as precipitating reagent
 - $Ca(OH)_2$ is used for testing carbon dioxide
4. Calcium Hydroxide used to soften temporary hard water.
5. Hydroxide of Aluminium and Magnesium uses as Antacids to neutralize stomach acid.

6. Calcium Hydroxide uses to make mortal.
7. Used in Bleaching of pulp. Calcium Hydroxide is used to prepare $(Ca(HSO_3)_2$) sulphite pulp (used in making paper and artificial silk) from wood.
8. Sodium Hydroxide used in making baking powder.

Quiz

Time: 7min

1. Mention two examples of
 - (a) Basic oxide (b) Amphoteric oxide
2. Briefly explain why aluminium oxide is not prepared by direct method?
3. Complete and balance the following equations



9.3 METAL CARBONATES

- These are the compounds which contains a metal and carbonate radical.
- They are formed when both hydrogen atoms in carbonic acid (H_2CO_3) are replaced by a metal.
- Aluminium Carbonate ($Al_2(CO_3)_3$) and Iron(II) Carbonate ($FeCO_3$) do not exist at all.

Preparation of Metal Carbonates

- There are two method used to prepare carbonates depending on whether the solubility of a particular metal carbonate.
 - i) Soluble carbonates
 - ii) Insoluble carbonates

i) Soluble carbonates

- Soluble carbonates prepared by reacts corresponding Alkali with Carbon dioxide.

- $NaOH_{(aq)} + CO_{2(g)} \rightarrow Na_2CO_{3(aq)}$
- $KOH_{(aq)} + CO_{2(g)} \rightarrow K_2CO_{3(aq)}$
- $NH_4OH_{(aq)} + CO_{2(g)} \rightarrow (NH_4)_2CO_{3(aq)}$

ii) Insoluble carbonates

- All Insoluble Metal Carbonate prepared by precipitation Method.
- Any soluble carbonate can be reacted with other soluble salt of the Metal to make carbonate.

- $MgCl_{2(aq)} + Na_2CO_{3(aq)} \rightarrow MgCO_{3(s)} + 2NaCl_{(aq)}$
- $Zn(NO_3)_{2(aq)} + K_2CO_{3(aq)} \rightarrow ZnCO_{3(s)} + 2KNO_{3(aq)}$
- $CaCl_{2(aq)} + K_2CO_{3(aq)} \rightarrow CaCO_{3(s)} + 2KCl_{(aq)}$

Classification of metal carbonates

◦ They classified into two due to Solubility in water, includes

Soluble Carbonates

◦ All Alkali metal carbonates like Sodium and Potassium are soluble in water.

◦ Exceptions: Lithium are slightly soluble.

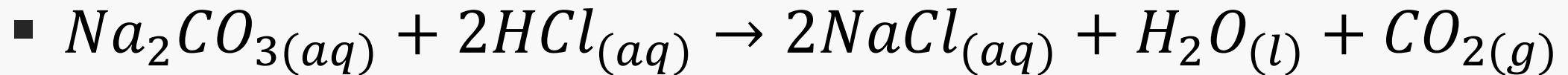
Insoluble Carbonates

◦ All left metal carbonates include of Calcium, Magnesium, Zinc, Iron and Copper are insoluble in water.

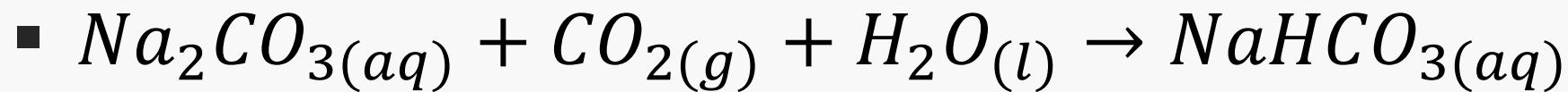
Properties of Metal Carbonates

1. **Solubility:** With exception to oxides of Na and K, all metal carbonates are insoluble(carbonate of **Li** is slightly soluble).
2. **Action of Heating:** Carbonates of Na & K do not decompose on heating but the rest decompose to give a metal oxide and Carbon dioxide.
 - $Na_2CO_{3(s)} \rightarrow No\ reaction$
 - $K_2CO_{3(s)} \rightarrow No\ reaction$
 - $MgCO_{3(s)} \rightarrow MgO_{(s)} + CO_{2(g)}$
 - $ZnCO_{3(s)} \rightarrow ZnO_{(s)} + CO_{2(g)}$

3. Action on Diluted Acids: Metal carbonates react with mineral acids to give a Salt, Carbon dioxide and Water.



4. Reacts with carbon dioxide in presence of water to form metal hydrogen carbonate.



Test for Carbonates

1. The soluble Carbonates when react with Magnesium Sulphate solution form white precipitates of Magnesium carbonates which is insoluble Salt.

- $Na_2CO_{3(aq)} + MgSO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + MgCO_{3(s)}$
- $K_2CO_{3(aq)} + MgSO_{4(aq)} \rightarrow K_2SO_{4(aq)} + MgCO_{3(s)}$

White precipitates

- This is the ***confirmatory test*** for *soluble* metal carbonates.

2. When metal carbonate reacts with dilute acid or decomposed produces CO_2 which turns lime water milky(white precipitates).

- $ZnCO_{3(s)} + HNO_{3(aq)} \rightarrow Zn(NO_3)_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$
 - $CaCO_{3(s)} + HNO_{3(aq)} \rightarrow Ca(NO_3)_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$
- The CO_2 produced can be tested using lime water
- $Ca(OH)_{2(aq)} + CO_{2(g)} \rightarrow CaCO_{3(s)}$

White precipitates

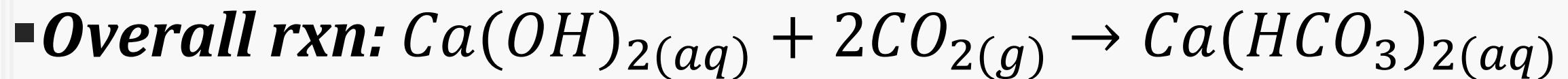
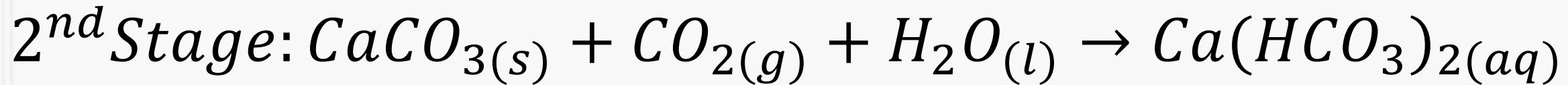
◦ This is the ***confirmatory test*** for *insoluble* metal carbonates.

9.4 METAL HYDROGEN CARBONATE

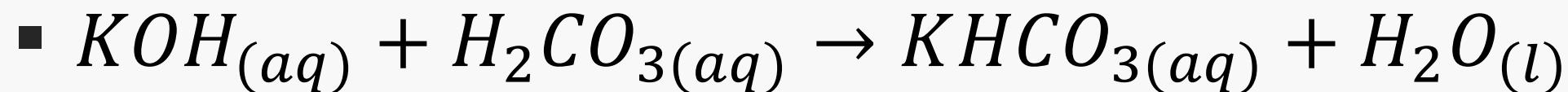
- These are the compounds which contains a metal and hydrogen carbonate radical.
- They are formed when one hydrogen atom in carbonic acid (H_2CO_3) is replaced by a metal.

Preparation of Metal Carbonates

- It is formed when excess carbon dioxide is passed through an aqueous solution of metal hydroxide.
- This involves two stages.



2. Replacement of one hydrogen atom from carbonic acid.



Properties of metal hydrogen carbonates

1. Solubility: All metal hydrogen carbonates are soluble in water.

2. Action of heat: All hydrogen carbonates decompose on heating to form metal carbonate, water and carbon dioxide.

- $Ca(HCO_3)_{2(s)} \rightarrow CaCO_{3(s)} + CO_{2(g)} + H_2O_{(l)}$
- $KHCO_{3(s)} \rightarrow K_2CO_{3(s)} + CO_{2(g)} + H_2O_{(l)}$

3. *Action of dilute acids:* Hydrogen carbonates reacts with dilute acids to form metal salt, water and carbon dioxide.

- $Ca(HCO_3)_{2(aq)} + HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$
- $KHCO_{3(aq)} + HNO_{3(aq)} \rightarrow KNO_{3(aq)} + CO_{2(g)} + H_2O_{(l)}$

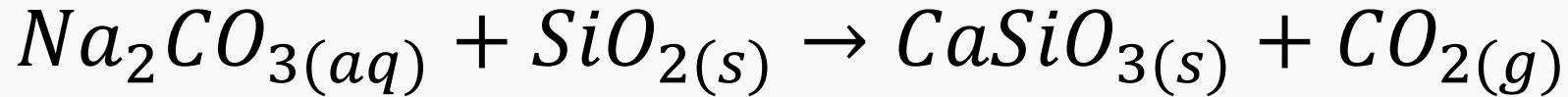
Test for metal hydrogen carbonates

- Metal hydrogen carbonates can be tested by using magnesium sulphate solution.
- When magnesium sulphate solution is added into a test tube that contain a metal hydrogen carbonate then heated, the white precipitates forms.
- $NaHCO_3(aq) + MgSO_4(aq) \rightarrow Na_2SO_4(aq) + Mg(HCO_3)_2(aq)$
- After heating $Mg(HCO_3)_2$ decomposes and form white precipitates
- $Mg(HCO_3)_2(aq) \rightarrow MgCO_3(s) + CO_2(g) + H_2O(l)$

White precipitates

Uses of metal carbonates and hydrogen carbonate

1. Sodium carbonate is used to soften hard water
2. Manufacture of glass: Glass is made by heating sand, calcium carbonate and sodium carbonate together, at a temperature of 1300°C to 1400°C.



◦Glass is a mixture of sodium and calcium silicates

3. Sodium Carbonates is used in the qualitative analysis of reactions involving carbonates because it is soluble in water.

4. Hydrogen carbonate of sodium is used to make baking powder.
5. Hydrogen carbonate of sodium is used to make anti-acid tablets.
6. Hydrogen carbonate of sodium is used as anti-fungal.

QUIZ

TIME: 05 Min

1. Mention three uses of metal carbonates.
2. A certain compound **Q** was heated but it didn't decompose. But when **Q** reacted with dilute hydrochloric acid gas **P** was produced, the gas **P** turned *lime water* milky. When compound **Q** was burned on a direct flame it burned with *golden-yellow* flame.
 - a) Identify the compound **Q** and gas **P**
 - b) Write the balanced chemical equation for the reaction between **Q** and dilute hydrochloric acid.

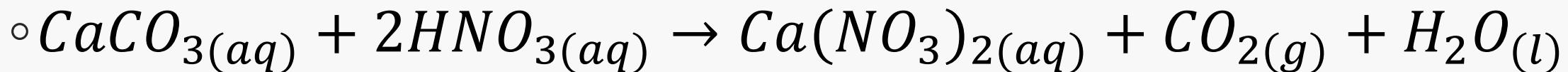
9.5 METAL NITRATES

- Metal Nitrates are the compounds that contain metal and nitrate radical.
- They are the salts delivered from Nitric Acid

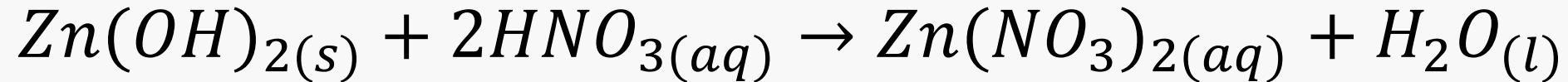
Preparation of metal nitrates

- Metal nitrates can be prepared by action of nitric acid on the following

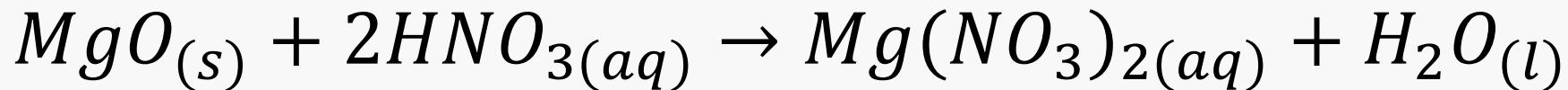
1. Metal carbonates



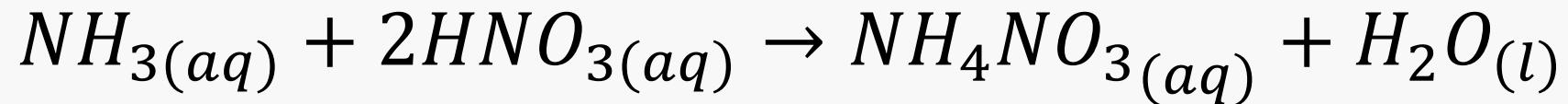
4. Metal hydroxide



5. Metal oxide.

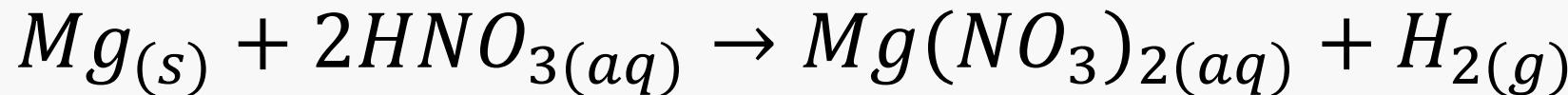


6. Ammonia solution.



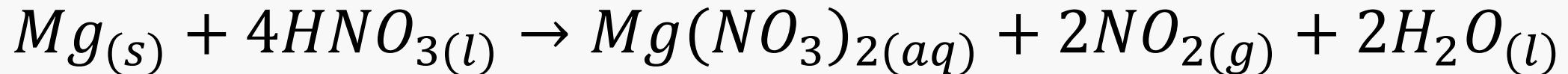
7. Metals: Here the product depends on the concentration of acid.

- A very dilute nitric acid(1%) reacts with magnesium to form magnesium nitrate and hydrogen.

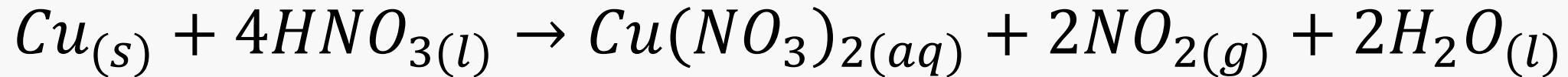


Diluted(1%)

- When the concentrated nitric acid reacts with magnesium to form magnesium nitrate, nitrogen dioxide and water



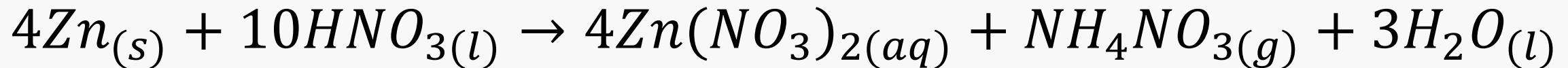
- When concentrated nitric acid reacts with copper to form copper nitrate, nitrogen dioxide and water



- When the acid is moderately concentrated, it reacts with copper to form copper nitrate, nitrogen monoxide and water



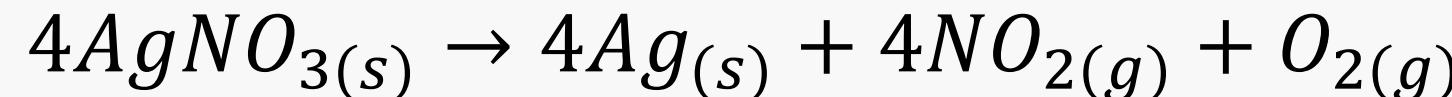
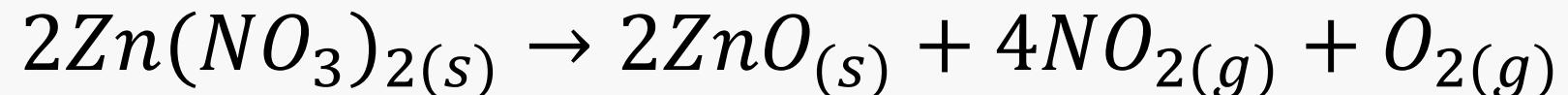
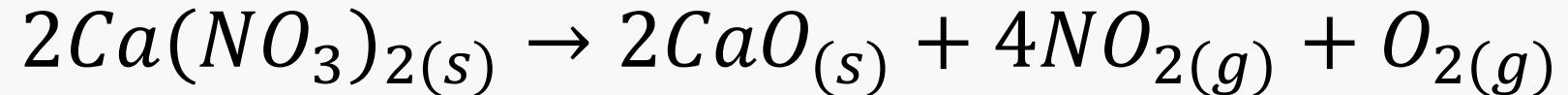
- Zinc reacts with to form zinc nitrate, ammonium nitrate and water.



Properties of metal nitrates

1. *Action of heat:* All metal nitrates decomposes on heating, but the products varies depending on the reactivity of a metal.

Potassium
Sodium
calcium
magnesium
aluminium
zinc
iron
lead
Copper
mercury
silver
gold
platinum



2. *Solubility:* All metal nitrates are soluble in water.

Test for nitrates

- The tests of nitrates include the following

Decomposition of metal nitrates

- With exception to nitrates of Na and K, metal nitrate decompose on heating and produce reddish brown fumes of nitrogen dioxide gas are observed.
- The gas produced(NO_2) signifies the presence of metal nitrates.
- $2Ca(NO_3)_2(s) \rightarrow 2CaO(s) + 4NO_2(g) + O_2(g)$



Saturday, September 26, 2020

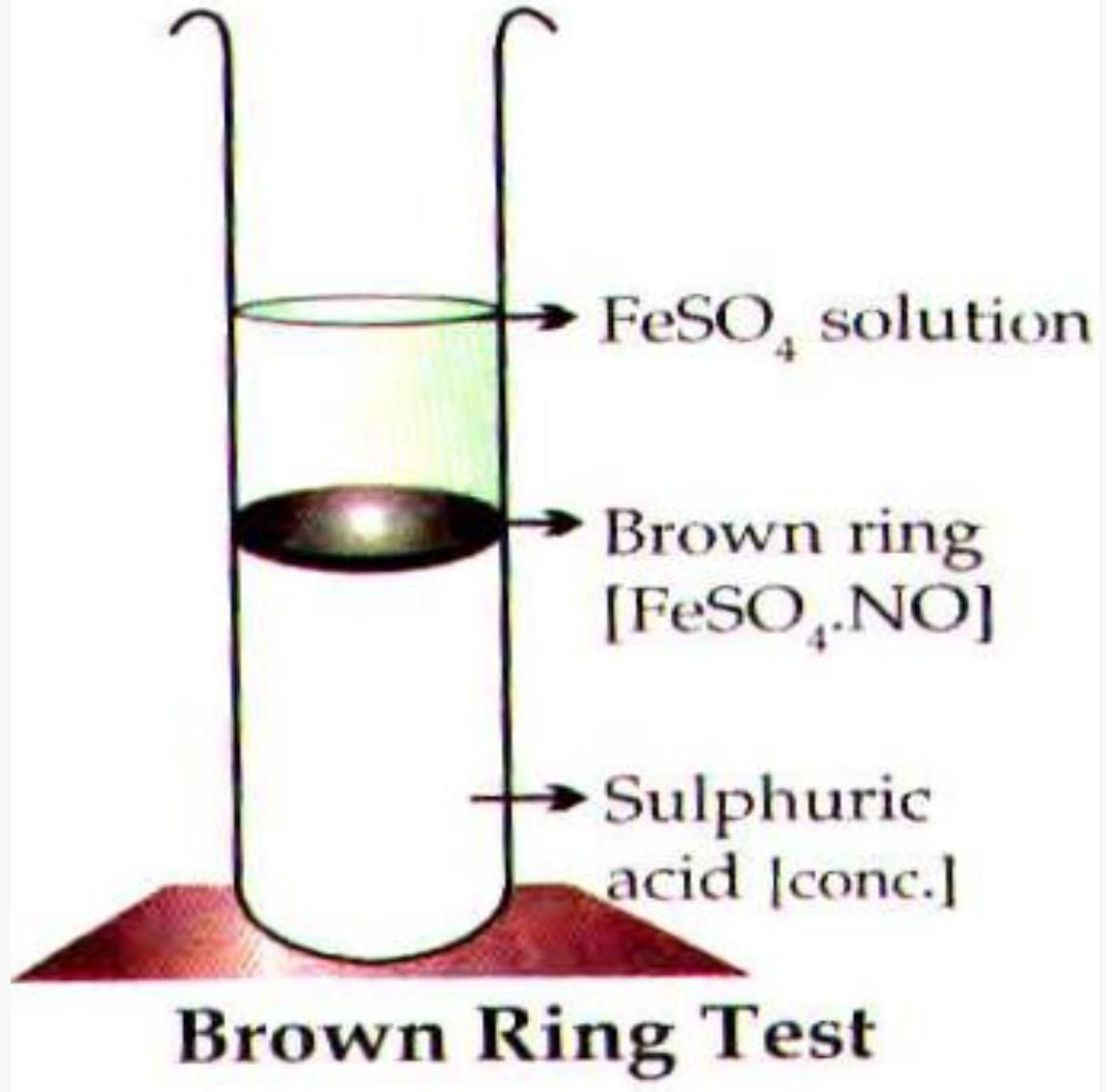


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Brown Ring Test

- A little amount of salt solution (about 2ml) is put in a test tube.
- Then a freshly prepared iron(ii) sulphate solution is added into a test tube containing sample solution.
- Finally the concentrated sulphuric acid is added down the side of the test tube carefully. The acid moves down at the bottom of the test tube as it is dense than the mixture.
- A brown layer forms at the junction between the two solutions.
- If the solution is disturbed(shaken) the ring disappears.



This is the *confirmatory test* for metal nitrates

Uses of metal nitrate

1. **Agriculture:** Nitrogen fertilizers are mainly Nitrate. They include Ammonium Nitrate, Potassium Nitrate, and calcium nitrate.
2. **Weapons:** Gun powder is a mixture of potassium nitrate, charcoal and sulphur. Ammonium nitrate is also used in making explosives and blasting agents which are used in mines and quarries.
3. **Food Preservation:** Nitrates and Nitrites are used in preserving meats and fish. Not only to kill Bacteria but also to produce a characteristic flavour.
4. Potassium nitrate is used in **salt bridge** in electrolysis.
5. **Antiseptics:** Is the one which used to kill Germs on the Skin.

9.6 METAL CHLORIDES

- These are the binary compounds which contains a metal and chlorine.
- Chlorides are salts that are derived from hydrochloric acid.
- Some of metals chlorides like common salts are very common in our daily life.

Preparation of metal chlorides

- There are three major methods of preparing metal chloride

Direct method

- This involves the reaction between metal and chlorine gas
- $2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$

Indirect method

- In this method metal chloride is prepared by the reaction between dilute hydrochloric acid and metal, metal oxide, metal nitrate, metal hydroxide, metal carbonate.
- $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$
- $CaO_{(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)}$
- $Cu(NO_3)_2_{(aq)} + 2HCl_{(aq)} \rightarrow CuCl_{2(aq)} + 2HNO_3_{(aq)}$

- $Zn(OH)_{2(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + 2H_2O_{(l)}$
- $MgCO_{3(aq)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

Precipitation method

- Insoluble chlorides are prepared by precipitation from their aqueous solutions.
- $Pb(NO_3)_{2(aq)} + 2HCl_{(aq)} \rightarrow PbCl_{2(s)} + 2HNO_{3(aq)}$
- $AgNO_{3(aq)} + HCl_{(aq)} \rightarrow AgCl_{(s)} + HNO_{3(aq)}$

Qn:

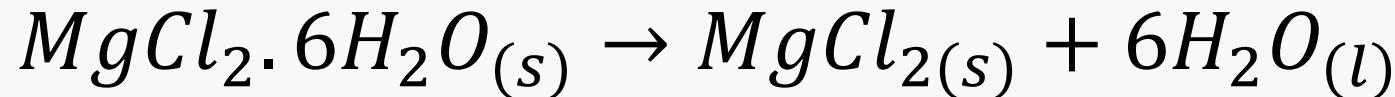
Why when HCl is added into lead(ii) oxide(or lead(ii) carbonate) the reaction starts then stop after a short time?

Ans:

- There may be a slight reaction but because lead(II) chloride is insoluble it forms on the surface of the lead(II) oxide(or carbonate) and inhibits further reaction.
- The standard way to prepare an insoluble salt is by ***precipitation***.
- You need to start with two suitable soluble reagents such as lead(II) nitrate and sodium chloride solution.

Properties of metal chlorides

1. Action of heat: All chlorides do not decompose on heat except ammonium chlorides. Hydrated magnesium chloride gives up its water of crystallization when it is gently heated



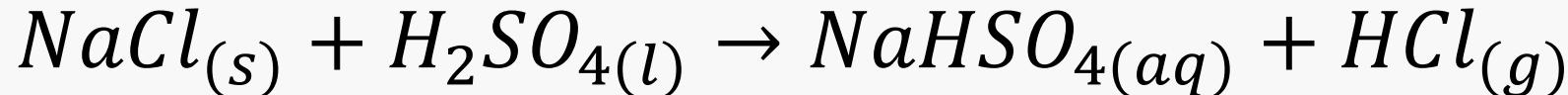
◦ When heated strongly it decompose to form a greenish-yellow gas(Cl_2), magnesium oxide and water.



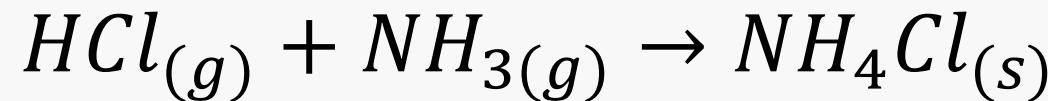
2. Solubility: All chlorides dissolves in water except chlorides of lead and silver.

Test for chlorides

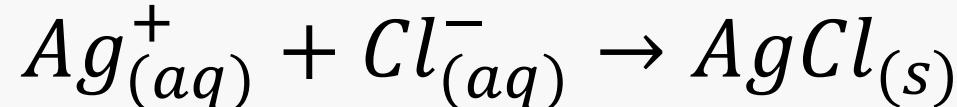
- When concentrated sulphuric acid is added to a metal chloride, hydrogen chloride gas is given off.



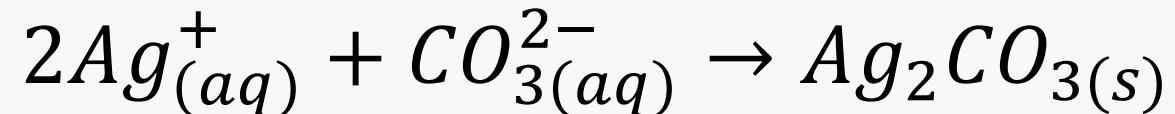
- Hydrogen chloride gas turns wet **blue** litmus paper into **red** and forms dense white fumes with gaseous ammonia.



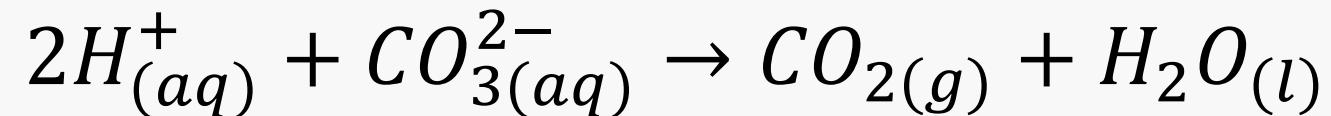
- Using silver nitrate solution: Few drops of dilute nitric acid are added first to a test solution followed by silver nitrate solution. White precipitate of silver chloride is formed.
- This is one of the ***confirmatory test*** for chlorides.



- Dilute nitric acid is added to prevent precipitation of other silver salts such as a carbonate, otherwise it will be mistaken for a chloride



- If a carbonate is present, nitric acid will remove it in form of carbon dioxide.



- Chlorides are not affected by the dilute nitric acid, but dissolve in ammonia solution.

Uses of metal chlorides

1. ***Manufacture of dry batteries:*** Zinc chloride mixed with ammonium chloride is used in dry batteries.
2. ***Household use:*** Common salt (sodium chloride) is added to food to give it taste.
3. ***Food preservation:*** Canned foods contain sodium chloride which in this case is used as a preservative to prevent the growth of bacteria.
4. ***Petroleum industry:*** Aluminium chloride acts as a catalyst in many organic reactions whose reactants are derived from petroleum.

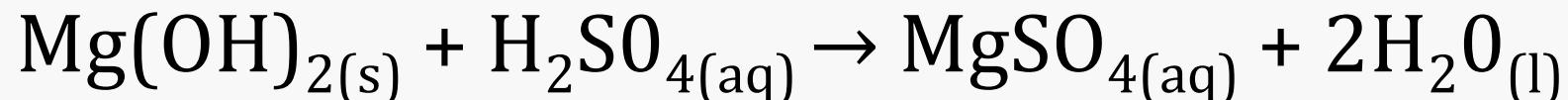
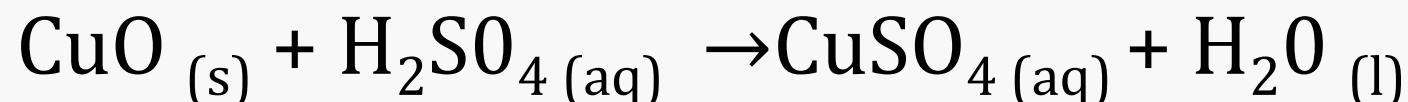
5. Aluminium chloride used in the manufacture of synthetic lubricating oils.
6. ***Spinning industry:*** Magnesium chloride is used to lubricate cotton threads in the spinning industry
7. ***Dentistry:*** Magnesium chloride is a constituent of the cement used to fill cavities in teeth.

9.7 METAL SULPHATES

- Metal sulphate is a salt that contains metal and sulphate radical.
- They are normally derived from sulphuric acid.

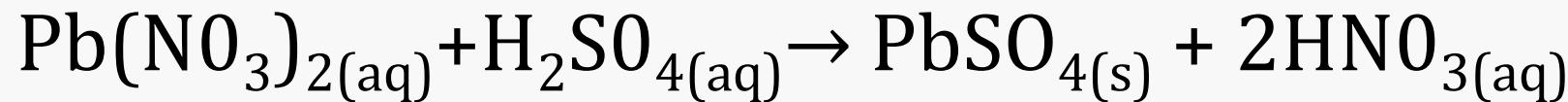
Preparation of metal sulphates

i) **Soluble sulphates** can be prepared by reacting a *metal, metal carbonate, metal hydroxide or metal oxide* with dilute sulphuric acid.

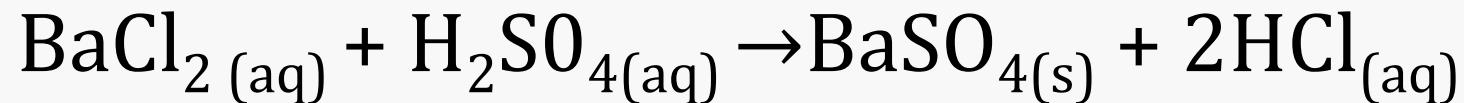


ii) Insoluble sulphates are prepared by precipitation reaction.

◦ Barium and lead sulphate are the only known insoluble sulphates (Calcium sulphate is slightly soluble).



White ppt



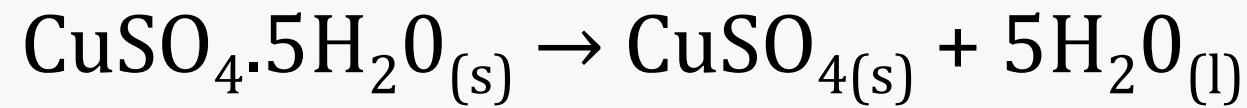
White ppt

Properties Of Sulphates

1) Solubility : With exception to Pb & Ba all sulphates are soluble in water(Calcium sulphate is slightly soluble).

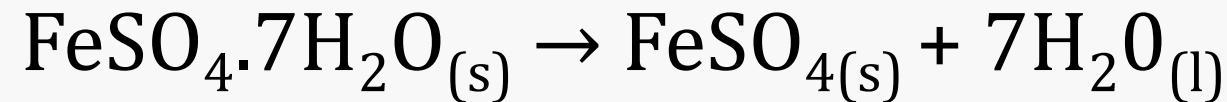
2) Action of heat : Sulphates are more stable to heat than the corresponding nitrates.

◦ Hydrated sulphates lose water when heated gently



Blue

White



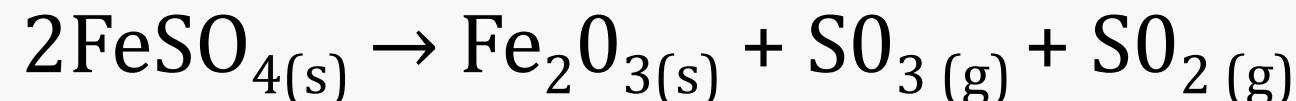
Light green

White

- But sulphates of metals low in the reactivity series must be strongly heated to decompose.
- Sulphates of Zn and Cu decompose to oxide and sulphur trioxide when heated strongly.



- On strong heating, the anhydrous iron(II) sulphate decomposes to form a reddish-brown iron(III) oxide, Sulphur trioxide and sulphur dioxide



- On strong heating, the anhydrous iron (III) sulphate decomposes to a reddish-brown iron (III) oxide and sulphur trioxide



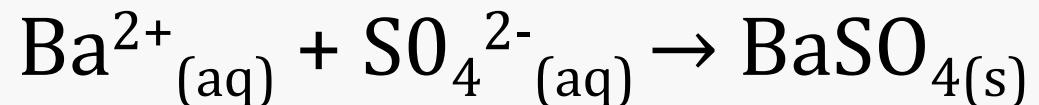
NB: The gases evolved when sulphates are heated strongly (SO_3 and SO_2) are acidic, hence they can turn the moist blue litmus into red.

Test for Sulphates

- Barium ions (Ba^{2+}) and lead ions (Pb^{2+}) are used to test for the presence of sulphate ions, but $\text{Ba}(\text{NO}_3)_2$ or BaCl_2 are mostly used.

How to test?

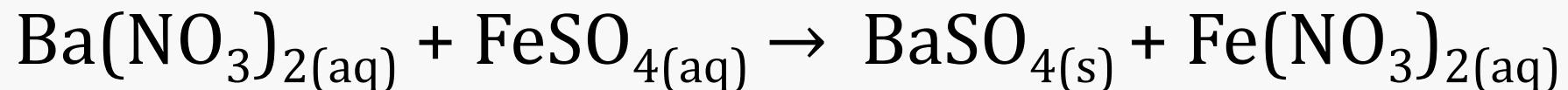
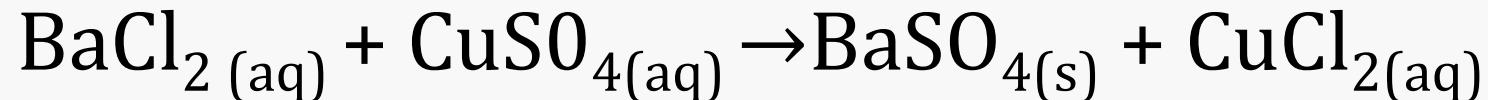
- To a sample solution few drops of dil. HCl are added followed by BaCl_2 . If white precipitates forms it signifies that SO_4^{2-} ions are present(dil HNO_3 and $\text{Ba}(\text{NO}_3)_2$ can be used the same way).



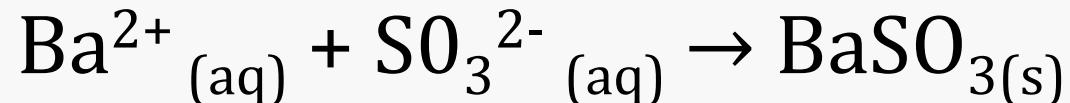
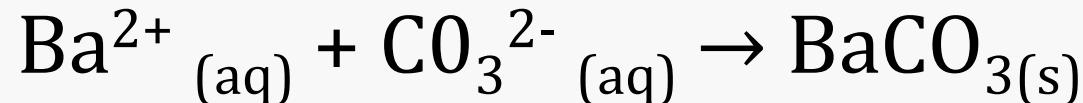
White precipitates

- This is the ***confirmatory test*** for sulphates.

Example:



- The acid is added to eliminate any **sulphites** and **carbonates**, which if present would also form precipitates with barium ions.



Uses of Sulphates

- 1) CaSO_4 is used to make Plaster of Paris (POP).
- 2) FeSO_4 is used to make tablets prescribed to patients who have iron deficiency.
- 3) FeSO_4 used as a weed killer.
- 4) FeSO_4 is used in qualitative analysis in test for nitrates.
- 5) FeSO_4 used for treating sewage and water
- 6) FeSO_4 used as a fungicide
- 7) BaSO_4 is used as white pigment in white paints.
- 8) BaSO_4 is used in taking X-ray pictures of the digestive tract

- 9) $\text{Al}_2(\text{SO}_4)_3$ is used as a size in paper making, it reduces the tendency of a paper to absorb liquid.
- 10) Hydrated zinc sulphate ($\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$) is used in wood preservation.
- 11) Hydrated zinc sulphate ($\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$) dyeing industry as an antiseptic.
- 12) Hydrated zinc sulphate ($\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$) is used in zinc-plating by electrolysis.

The end

Thank you for

*Thank you for
Your cooperation ..!!*