# CHEMICAL EQUATION

A chemical equation is a short way of describing exactly what happens during a chemical reaction.

Chemical equations must;

1. Tell the truth or represent the facts

2. Show symbols of elements and formula of products and reactants

3. Be balanced

4. Bear state symbols (s), (l), (g) and (aq) for substances dissolved in water (aqueous – water).

A state symbols s, l, g and aq that appear in chemical equation described as;-
s - Solid state of chemical component.
l - Liquid state of chemical component.
g - gaseous state of chemical component.
aq - a dissolved chemical component in water.

NB: Reactants and products are separated using an arrow $\rightarrow$

$\rightleftharpoons$ If the reaction is reversible

Example; $NH_4Cl_{(s)} \rightleftharpoons NH_{3(g)} + HCl_{(g)}$

$H_2 + 0_2 \rightleftharpoons H_2O$

## BALANCING OF CHEMICAL EQUATION

Balancing of the chemical equations is done so as to obey the law of conversation of matter which states;

"Matter can never be created nor destroyed in a chemical reaction ".

When balancing the chemical equation we change the coefficient only.

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Examples;

1. $O_2 + H_2O \rightarrow H_2O$

Consider no of atoms

$O_2 + 2H_2 \rightarrow 2H_2O$

right side 6 atoms
left side 6 atoms

The coefficient of chemical symbol describe single molecule of component. A component of single molecule, its coefficient does not written on chemical symbol and considered as single molecule e.g $O_2$ in the equation above. And all unit multiple coefficients are written on chemical symbol e.g 2H2 in equation above.

2. $Cl_2 + 2KI \rightarrow 2KCl + I_2$

3. $AlCl_3 + 3NaOH \rightarrow Al(OH)_3 + 3NaCl$

4. $2KClO_3 \rightarrow 2KCl + 3O_2$

- When balancing don't consider anything which is above or below the line

- When balancing statement or equations like this find the 1 cm of the two numbers 2, 3 to get 6

5. $KNO_3 \rightarrow KNO_2 + O_2$

- If oxygen is alone then reduce the oxygen atoms

$KNO_3 \rightarrow KNO_2 + {}^{1}/_{2}\, O_2$

- Eliminate denominator by multiplying 2 throughout

$2KNO_3 \rightarrow 2KNO_2 + O_2$

6. $Cu(NO_3)_2 \rightarrow CuO + NO_2 + O_2$

$Cu(NO_3)_2 \rightarrow CuO + 2NO_2 + {}^{1}/_{2}\,O_2$

$2Cu(NO_3)_2 \rightarrow 2CuO + 2NO_2 + O_2$

$2Cu(NO_3)_2 \rightarrow 2CuO + 4NO_2 + O_2$

7. $NH_4NO_3 \rightarrow N_2O + H_2O$

$NH_4NO_3 \rightarrow N_2O + 2H_2O$

8. $NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$

$2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$

9. $AgNO_3 \rightarrow Ag + NO_2 + O_2$

$2AgNO_3 \rightarrow 2Ag + 2NO_2 + O_2$

10. $Al(OH)_3 \rightarrow Al_2O_3 + H_2O$

$2Al(OH)_3 \rightarrow Al_2O_3 + 3H_2O$

11. $C_2H_6 + O_2 \rightarrow H_2O + CO_2$

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

$2C_2H_6 + {}^{7}/_{2}\, x\, 2O_2 \rightarrow 2x3H_2O + 2x2CO_2$

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

**Types of chemical reactions include,**

## 1. Decomposition reactions.

Decomposition reaction is that reaction, where by a giant molecules of chemical component break-up into simple molecules. Decomposition can be influenced by heating or catalysis.

Examples; **By heating**

$$CaCO_{3(s)} \xrightarrow[\Delta]{heat} CaO_{(s)} + CO_{2(g)}$$

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**By catalyst**

$$2H_2O_{2(L)} \xrightarrow[\triangle]{KMnO_4} 2H_2O_{(L)} + O_{2(g)}$$

There are two types of decomposition by heating.

**i) Thermal decomposition.**

Is the type of decomposition reactions which are irreversible.

**Example;**

$$CaCO_{3(s)} \xrightarrow[\triangle]{heat} CaO_{(s)} + CO_{2(g)}$$

**NOTE;**

$\triangle$ is a symbol for decomposition

**(ii) Thermal dissociation**

Is a type of decomposition reaction which are reversible.

Example; $NO_2O_4 \rightleftharpoons 2NO_2$

$$NH_4Cl_{(s)} \rightleftharpoons NH_{3(g)} + HCl_{(g)}$$

2. **Synthesis (combination) reactions**

Simpler compounds combine to form one compound (opposite of decomposition)

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

$$H_2 + I_2 \rightarrow 2HI$$

3. **Displacement reactions**

In these reactions one element displaces another element from a compound

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Example; $Zn_{(s)} + HCl_{(aq)} \rightarrow ZnCl_{(aq)} + H_{2(g)}$

$Cl_{2(g)} + KI_{(aq)} \rightarrow KCl_{(aq)} + I_{2(s)}$

4. **Double decomposition reactions**

The compound exchange their radicals

$AB + CD \rightarrow AD + CB$

B and D are only to be interchanged. The reactants are always aqueous

$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

The reactant must be soluble in water and one of the product must be either solid or a gas or liquid

5. **Neutralization**.

This is the type of reaction which involves acid and base which react to produce salt and water only. All metal oxide is basic in nature.

Base + Water = Alkali. Not all bases are alkalis but all alkalis are bases because not all bases form alkalis.

Examples,

$CaO + HCl \rightarrow H_2O + CaCl_2$

$CaO + 2HCl \rightarrow H_2O + CaCl_2$

Example;

$NaOH + H_2SO_4 \rightarrow H_2O + Na_2SO_4$

$2NaOH + H_2SO_4 \rightarrow 2H_2O + Na_2SO_4$

**Question;**

Complete and balance;

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Classify the following chemical reactions;

1. $2Mg + O_2 \rightarrow 2MgO$

   Synthesis reaction

2. $N_2 + 3H_2 \rightarrow 2NH_3$

   Synthesis reaction

3. $2H_2O_2 \rightarrow 2H_2O + O_2$

   Decomposition reaction

4. $CuSO_2 + H_2S \rightarrow CuS + H_2SO_4$

   Double decomposition

5. $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

   Displacement reaction

Writing a balanced chemical equation for the following:

I. Iron + Sulphur → Iron sulphide

II. Decomposition of calcium carbonate

III. Magnesium hydroxide + Hydrochloric acid → Magnesium chloride + Sodium chloride

IV. Silver nitrate + Sodium chloride → Silver chloride+ Sodium nitrate

V. Potassium + Sulphuric acid → Potassium sulphate + Hydrogen gas

## IONIC EQUATIONS

These are reactions equations written by omitting the spectator ions (ions remaining even after the reaction.

## STEPS OF WRITING IONIC EQUATION

1. Balance the equation;

$$2HCl_{(aq)} + CaCO3_{3(s)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + HO_{2(l)}$$

2. Split all solid and soluble ionic compound on the reactants side and only the soluble ionic compounds on the product side into individual ions

$2H^{+}{}_{(aq)} + 2Cl^{-}{}_{(aq)} + Ca^{2+}{}_{(s)} + CO^{2-}{}_{(aq)} \rightarrow Ca^{2+}{}_{(aq)} + 2Cl^{-}{}_{(aq)} + CO_{2(g)} + HO_{2(l)}$

3. Omit the ions found on both side of the equation

$2H^{+}{}_{(aq)} + 2Cl^{-}{}_{(aq)} + Ca^{2+}{}_{(s)} + CO^{2-}{}_{(aq)} \rightarrow Ca^{2+}{}_{(aq)} + 2Cl^{-}{}_{(aq)} + CO_{2(g)} + HO_{2(l)}$

4. Write the equation for remaining species (ionic equation)

$2H^{+}{}_{(aq)} + CO^{2-}{}_{(aq)} \rightarrow CO_{2(g)} + HO_{2(l)}$

**QUESTION;**

1. $NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

$Na^{+}{}_{(aq)} + Cl^{-}{}_{(aq)} + Ag^{+}{}_{(aq)} + 3NO^{-}{}_{(aq)} \rightarrow AgCl_{(s)} + Na^{+}{}_{(aq)} + 3NO^{-}{}_{(aq)}$

$Cl^{-}{}_{(aq)} + Ag^{+}{}_{(aq)} \rightarrow AgCl_{(s)}$

2. $BaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow BaSO_{4(aq)} + 2NaCL_{(aq)}$

$Ba^{2+}{}_{(aq)} + 2CL^{-}{}_{(aq)} + 2Na^{+}{}_{(aq)} + SO^{2-}{}_{4(aq)} \rightarrow BaSO_{4(aq)} + 2Na^{+}{}_{(aq)} + 2Cl^{-}{}_{(aq)}$

$Ba^{2+}{}_{(aq)} + SO^{2-}{}_{4(aq)} \rightarrow BaSO_{4(aq)}$

3. $NaCl_{(aq)} + H_2SO_{4(aq)} \rightarrow NaHSO_{4(aq)} + HCL_{(aq)}$

$Na^{+}{}_{(aq)} + CL^{-}{}_{(aq)} + H^{+}{}_{(aq)} + HSO^{-}{}_{4(aq)} \rightarrow NaHSO_{4(s)} + H^{+}{}_{(aq)} + Cl^{-}{}_{(aq)}$

$Na^{+}{}_{(aq)} + HSO^{-}{}_{4(aq)} \rightarrow NaHSO_{4(s)}$

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## ACIDS, BASES AND SALT

**BASE**

A base is a substance which contains the oxide ion or hydroxide ion and reacts with acid to give salt and water only

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

$CaO + HCl \rightarrow CaCl_2 + H_2O$

Or

A base is a soluble substance which accepts a proton

Example; $H^+ + NH_3 \rightarrow NH_4^+$

A soluble base is called Alkali

The Alkali's are

Potassium hydroxide → KOH

Sodium hydroxide → Na OH

Ammonium hydroxide (Ammonia solution) = $NH_4OH$

Calcium hydroxide (slacked line) → $Ca(OH)_2$

The solution of $Ca(OH)_2$ in water is called lime water

$Ca(OH)_2$ = Sparingly soluble

Magnesium hydroxide (milk of magnesia ………$Mg(OH)_2$

It's also sparingly soluble

The soluble bases when in solution produce OH- as their only negative ion

Example; $NaOH_{(aq)} \rightarrow Na^+ + OH^-$

$NH^+{}_4 + OH^- \rightarrow NH_4OH_{(aq)}$

## PROPERTIES OF BASES

1. Alkalis have a soapy feeling (slippery)

2. Alkalis have bitter taste

3. Alkalis turn red to blue litmus paper

4. They react with acids and water only

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Question: which bases combine to form NaCl?

5. Alkalis yield ammonia when warmed with ammonium salts

Example; $NH_4Cl + KOH \rightarrow NH_3 + H_2O + KCl$

$(NH_4)_2SO_4 + NaOH \rightarrow NH_3 + H_2O + Na_2SO_4$

Examples; $\text{Acid} + \text{Base} \rightarrow \text{Salt} + \text{Water}$

## USES OF BASES

a) Softening water. Example; Calcium hydroxide

b) Making fertilizers. Example; $(NH_4)_2SO_4$ as sulphate of Ammonia

c) Used as antacids in stomach. Example; Mg (OH), CaO

d) For manufacturing of paper soap etc especially in industries

## QUESTION:

Calculate percentage of hydrogen in compound (HCL)

H + Cl

1 + 35.5 = 36.5 (atomic masses)

H% = 1/36.5 x 100%

Cl% = 35.5/36.5 x 100%

What kind of equation is?

$NH_4Cl + KOH \rightarrow NH_3 + H_2O + KCl$

Answer: Special case reaction

## ACIDS

These are the substances which when in solution produce hydrogen ions as their only positive ions

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Example; $HCl_{(aq)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$

$H_2SO_{4(aq)} \rightarrow 2H^{+}_{(aq)} + SO^{2-}_{4(aq)}$

An acid is a substance which turns blue litmus to red and which contains hydrogen replaceable by a metal

Example; $Zn_{(s)} + HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

$Ca_{(s)} + H_2SO_{4(aq)} \rightarrow CaSO_{4(aq)} + H_{2(g)}$

## TYPES OF ACIDS

There are two types of acids i) Mineral acids

ii) Organic acids

### A: Mineral Acids

These are derived from substances formed from minerals, for example; from minerals

| Acid | Formula | Source | Formula |
|---|---|---|---|
| Hydrochloric acid | $HCl$ | Common salt | $NaCl$ |
| Sulphuric acid | $H_2SO_4$ | Sulphur | $S$ |
| Nitric acid | $HNO_3$ | Chilean salt petre | $NaNO_3$ |
| Carbonic acid | $H_2CO_3$ | Limestone | $CaCO_3$ |

The mineral salts that can be derived from the oxides are called oxy - acids (from the acidic/ non – metal oxides)

### B. ORGANIC ACIDS

These are derived from plants and animals source animal source (organic sources)

Example;

| Acid | Source |
|---|---|

| | |
|---|---|
| Lactic acid | Sour milk |
| Acetic acid | Vinegar |
| Tartaric acids | Grapes |
| Citric acid | Citreous fruits |

## STRENGTH OF ACIDS

Strength of acids are classified into two i) Strong acids

ii) Weak acids

## STRONG ACIDS

These are acids which ionize completely in aqueous solution

Example; $HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

$$H_2SO_{4(aq)} \rightarrow 2H^+_{(aq)} + SO^{2-}_{4(aq)}$$

$$HNO_{3(aq)} \rightarrow H^+_{(aq)} + NO^-_{3(aq)}$$

## WEAK ACIDS

These are acids which ionize partially in aqueous solution

$$H_2CO_{3(aq)} \rightleftharpoons 2H^+_{(aq)} + CO^{2-}_{3(aq)}$$

$$CH_3COOH \rightleftharpoons CH_3COO_{(aq)} + H^+$$

All organic acids are weak

## QUESTION;

Is phosphoric acid a strong or weak acid? What are strong bases or weak bases?

## PROPERTIES OF ACIDS

1. Have sour taste

2. Action on indicators

Indicator is a substance which changes color in acids or bases

Example; litmus, methyl orange, phenolphthalein etc

| Indicators | Color in acids | Color in alkali |
|---|---|---|
| Methyl orange | Pink | Yellow |
| Phenolphthalein | Colorless | Pink |
| Bromothymol blue | Yellow | Blue |
| Litmus | Red | blue |

3. Corrosive action

A concentrated acid is a water molecule aspires. Acid dries any substance of water content in vigorous action, thus burns skin of animals, leaves and stems of plants.

Example; concentrated $H_2SO_4$, HCl etc

4. Action with metal

Any metal which is above hydrogen in the activity series will replace hydrogen from the acid. The products of these reactions are salt and hydrogen gas

Example; $Zn_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$

5. Action with metal oxides and hydroxides to form salt and water only

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Example; a) with metal oxides

$$ZnO + 2HCl \rightarrow ZnCl_2 + H_2O$$

$$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$$

b) With hydroxides

$$NaOH + HCl \rightarrow NaCl + H_2O$$

$$KOH + HNO_3 \rightarrow KNO_3 + H_2O$$

6. Action with carbonates and hydrogen carbonates (bicarbonates)

Acids react with carbonates and bicarbonates producing salt, water and $CO_2$

Example; $CaCO_{3(s)} + 2HNO_{3(aq)} \rightarrow Ca(NO_3)_2 + H_2O_{(l)} + CO_{2(g)}$

$$PbCO_{3(s)} + 2HCl_{3(aq)} \rightarrow PbCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

7. Action with sulphates

When acids react with sulphates, $SO_2$, water and salts are produced

Example; $Na_2SO_3 + 2HCl \rightarrow 2NaCl + H_2O + SO_2$

## **BASICITY OF ACIDS**

The basicity of an acid is the numbers of hydrogen ions ($H^+$) produced by one mole of an acid in aqueous solution

Example; $HCl \rightarrow H^+ + Cl^-$

monobasic

$$H_2SO_4 \rightarrow 2H^+ + H_2O + SO_4^{2-}$$

Dibasic

$$H_3PO_4 \rightarrow 3H^+ + PO_4^{3-}$$

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Tribasic

$CH_3COOH$

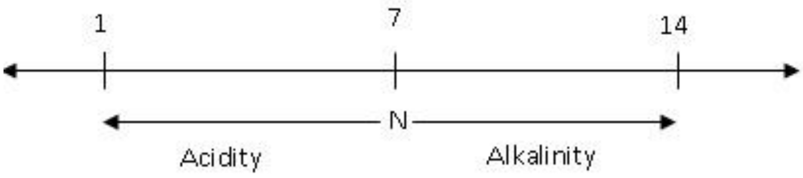
Monobasic

**NB:** All organic acids are monobasic

**$P^H$ SCALE;-**

It isn't possible to know exactly how strong an acid or base is by using an indicators. The term PH is used to describe or show the exact strength of an acid or base. It shows the number of hydrogen ion concentration

The $P^H$ scale can be represented diagrammatically as show below;



1-4 Strong acid

5-6 Weak acid

7 Neutral

8-9 Weak alkali

10-14 Strong alkali

**USES OF ACIDS**

Acids have the following uses;

a) Preparation of salts, for example; from fertilizers

From bases / metals.
$Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$

$NH_4OH + H_2SO_4 \rightarrow (NH_4)_2SO_4 + 2H_2O$

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b) Preparation of other acids;

e.g.$Pb(NO_3)_2 + H_2SO_4 \rightarrow PbSO_4 + 2HNO_3$

**Example;**

c) Manufacture of artificial silk

d) Clearing of metals, acids can combine with them to form salts can be easily removed out

e) Necessary for digestion of protein in the stomach.

## SALTS

Definition: A salt is a compound consisting of a positive metallic ion (or) and negative ion derived from an acid

Types of salts;

I) Acidic salts

ii) Normal salts

iii) Basic salts

### 1. Acidic salts

This is the salt which contains part of the hydrogen ions of an acid

Example; Sodium hydrogen – sulphate

Calcium hydrogen carbonate

### 2. Normal salts

These are the salts which contain neither $H^+$ ions from the acid nor $O^{2-}$ or $OH^-$ ions from the base

Example; $NaCl$, $MgCl_2$,$MgSO_4$ ,$FeCl_2$ ,$CaCO_3$

### 3. Basic salts

These are salts containing or ions from the base

Example; MgOHCl, ZnOHCl

Magnesium hydroxyl chloride and zinc hydroxyl chloride

O is incorrect because oxygen exists in molecule state but not in atomic

In previous salts was classified according to the nature of salts but in terms of solubility there are two types of salts (i ) Soluble salt ( ii ) Insoluble salt.

**SOLUBILITY RULE** states;

**Soluble** salts are;-

a) All salts of (made of) K, Na and $NH^+_4$

Example; $KNO_3$,KCl ,$K_2SO_4$ ,$K_3PO_4$ ,$K_2CO_3$,$NaHCO_3$

This is an acidic salt (because it gives out 2 positive ions of Na and H)

1. All nitrates

   Example; $KNO_3$

ii) All hydrogen carbonates

Example; $NaHCO_3$

b) All chlorides except

i) Silver chloride (AgCl)

ii) Mercury (I) Chloride (HgCl) iii) Lead (II) chloride ($PbCl_2$) which is soluble in hot water

N.B: In reactions the compound above are considered as solids (s)

c) All sulphates except

i) - Barium sulphate.

ii) - Lead (II) sulphate.

iii) - Calcium sulphate is slightly soluble in water.

**Insoluble salt** are;

a) All carbonates except those of K, Na, $NH_4$

b) All hydroxyl except those of K, Na, and Ca (which is slightly soluble in water)

Salt of strong acid and bases do not react with water

**Solubility rule** (summary)

i) All salts of K, Na and are soluble

ii) All nitrates are soluble

iii) All sulphates are soluble except those Ba, lead (II) and Ca

Vi) All chlorides are soluble except those Ag, lead (ii) and mercury (x)

All carbonates and hydroxides are insoluble except those of K, Na,

**NB:** - All acids are soluble unless when concentrated (aq)

- All oxides are labeled by solids (s)

**Examples;**

## METHODS OF PREPARING SALTS

### A: SOLUBLE SALTS

i) Direct combination of elements

Example; Burning Mg in chlorine gas

$$Mg + Cl_2 \rightarrow MgCl_2$$

Pass hot Al in Chlorine gas.

$$2Al_{(hot)} + 3Cl \rightarrow 2AlCl_3$$

ii) Addition of a metal to a dilute acid. But this method is restricted to metals above hydrogen in the reactivity series.

A metal below hydrogen cannot displace hydrogen

Example;

$Zn + 2HCl_2 + H_{2(g)}$

$Na + H_2SO_4 \rightarrow Na_2SO_4 + H_{2(g)}$

1. Addition of carbonates to a dilute acid.

   Products = salt, water and carbon dioxide.

   $CaCO_{3(s)} + 2HCl \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

2. Addition of an oxide to a dilute acid.

   $CaO + H_2SO_4 \rightarrow CaSO_4 + H_2O_{(l)} + H_2O$

v) The general method for preparing soluble salts is called crystallization, for example; Preparation of $ZnSO_4$

**Procedure**;

a) Add excess Zn to dilute in a beaker Add few crystals of and few drops of concentrated to speed up the reaction rate

1. If the reaction is still slow warm the mixture
2. When the reaction is over, filter and put the filtrate in an evaporating dish

c) Crystallization

The filtrate is evaporated and cooled. Check if the crystals are formed. If they are formed stop evaporating and leave it for clay.

The water attached or associated to the crystals is called **water of crystallization.**

$ZnSO_4.7H_2O$ is called zinc sulphate hepta hydrate.

B. **INSOLUBLE SALTS**

The insoluble salts are prepared by ionic precipitation also called double decomposition

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Here two soluble salts are mixed and react by interchanging their radicals ion forming both soluble and insoluble salts.

$BaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2NaCl$

$Ba^{2+}{}_{(aq)} \cancel{+ 2Cl}_{(aq)} + \cancel{2Na}^{+}{}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)} + 2\cancel{Na}^{+}{}_{(aq)} \cancel{+ 2Cl}^{-}{}_{(aq)}$

$Ba^{2+}{}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)}$

## QUESTION;

Ions to get solid, what is the name of the reaction? It's a double decomposition reaction or ionic precipitation.

## PROPERTIES OF SALTS

### 1. COLOUR

Some salts are coloured

For example; i) Iron (ii) salts are green

ii) Iron (ii) salts are yellow

iii) Iron (ii) salts are pale blue

iv) Nickel (ii) salts are green

v) Cobalt (ii) salts are pink

The colours of the above mentioned salts are due to the color of hydrated ion

Coloured salts are formed only in the group of transition elements. All elements in first 20 elements are not colored.

Other salts are not colored, for example; salts of Na, K, Ca, Pb, Zn, Al, Mg

### 2. Hydrolysis

This is the reaction of salts and water giving an acid or alkaline solution

Example:

$NH_4Cl_{(aq)} + H_2O \rightarrow NH_4OH_{(aq)} + HCl_{(aq)}$

$Na_2CO_{3(aq)} + H_2O_{(L)} \rightarrow NaOH_{(aq)} + H_2CO_{3(aq)}$

a) All salts of weak bases and strong acids hydrolyze to give acid solution

Example; $NH_4$, Cl, $FeCl_2$, $CuCl_2$,$Al_2(SO_4)$ etc.

b) All salts of strong bases and weak acids are hydrolyzed to give an alkaline solution

Example; $Na_2CO_3$ ,$CH_3COONa$, etc.

**N.B:** Salts of strong bases and strong acids are not undergoing hydrolysis. They only ionize when in solution:

Example; NaCl + $H_2O$ → There's no reaction.

## 3. EXPOSURE TO AIR

When salts are exposed to air they either lose water of crystallization or absorb water from the atmosphere

### a) Hygroscopic

Is the action of absorbing water from the atmosphere without changing into solution NaCl, anhydrous copper II sulphate ($CuSO_4$) and $NaNO_3$ others are concentrated and copper II sulphate

### b) Deliquescent

This is the action of absorbing water from the atmosphere by a solid to form a solution

Example; $MgCl_2$, $CaCl_2$, $FeCl_3$, $Ca(NO_3)_2$

Others are the hydroxides of K and Na

### c) Fluorescence

This is the action of giving up water of crystallization of the solid to the atmosphere

**Examples;**

i) Hydrate sodium carbonate

$Na_2CO_3.10H_2O \rightarrow Na_2CO_3.H_2O + 9H_2O$

ii) Hydrated sodium sulphate

$Na_2SO_4.10H_2O \rightarrow Na_2SO_4 + 10H_2O$

iii) Hydrated magnesium sulphate

$MgSO_4.7H_2O \rightarrow MgSO_4.H_2O + 6H_2O$

## 4. HEAT EFFECTS ON SALTS

Different salts behave differently on heating. Most of the hydrated salts loose water of crystallization when heated. The anhydrous salts undergo chemical change when heated

### I. Sulphates

a) Hydrated copper (ii) sulphate [$CuSO_4.5H_2O$]– changes colour from blue to white when heated

$$\underset{\text{Blue}}{CuSO_4.5H_2O} \xrightarrow[\Delta]{} \underset{\text{white}}{CuSO_4 + 5H_2O_{(g)}}$$

- If white copper (ii) sulphate ($CuSO_4$) is heated more strongly it becomes to black copper (ii) oxide (Cu

$$\underset{\text{White}}{CuSO_4} \xrightarrow[\Delta]{} \underset{\text{Black}}{CuO_{(S)} + SO_{3(g)}}$$

b) Hydrated iron (ii) sulphate loose all its water of crystallization on heating and changes from green to white

$$\underset{\text{Green}}{FeSO_4.7H_2O_{(s)}} \xrightarrow[\Delta]{} \underset{\text{white}}{FeSO_{4(s)} + 7H_2O_{(g)}}$$

c) When iron (ii) sulphate is heated strongly it decomposes to form black iron (iii) oxide, and $SO_3$ and $SO_{2(g)}$

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$2FeSO_4 \rightarrow Fe_2O_{3(s)} + SO_{3(g)} + SO_{2(g)}$
White black

d) Iron (iii) sulphate when heated strongly it decomposes to form black iron (iii) oxide and $SO_3$ only.

$Fe_2(SO_4)_{3(s)} \rightarrow Fe_2O_{3(s)} + 3SO_{3(g)}$

White black

**II) Chlorides**

All chlorides of metals are hydrated except those of K, Na, Pb, Hg and Ag when heated, chlorides undergo a chemical change called hydrolysis (i.e. they don't form anhydrous chlorides) in which hydrogen chloride gas (and water involve and a basic salts of chloride or oxide are formed

**Example;**

$$MgCl_2.6H_2O_{(s)} \xrightarrow[\triangle]{} MgOHCl_{(s)} + HCl_{(g)} + 5H_2O_{(g)}$$

$$AlCl_3.6H_2O_{(s)} \xrightarrow[\triangle]{} AlO_{3(s)} + 8H_2O_{(g)} + 6HCl_{(g)}$$

$$ZnCl_2.6H_2O_{(s)} \xrightarrow[\triangle]{} ZnOHCl_{(s)} + HCl + 5H_2O_{(g)}$$

Ammonium chloride sublimes when heated

$$NH_4Cl_{(s)} \xrightarrow[\triangle]{} NH_{3(g)} + HCl_{(g)}$$

**III) Carbonates and hydrogen carbonates**

Carbonates of sodium and potassium are unaffected by heat (even at very high temperature)

Ammonium carbonate decomposes readily on heating to form $NH_{3(g)}$, $CO_{2(g)}$ and $H_2O_{(g)}$

$$(NH_4)_2CO3_{(s)} \xrightarrow[\triangle]{} 2NH_{3(g)} + CO_{2(g)} + H_20_{(g)}$$

All other carbonates decompose to give oxide and $CO_2$

$$Al_2(CO_3)_3 \xrightarrow[\triangle]{} Al_2O_3 + CO_2$$

$$K2CO3 \xrightarrow[\triangle]{} \text{NO Reaction}$$

All hydrogen carbonates decompose to give metal carbonates water and $CO_2$ on heating

$$Mg(HCO_3)_2 \rightarrow MgCO_3 + H_2O + CO_2$$

$$2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$$

**IV) Nitrates**

Na and K nitrates decompose when heated to give corresponding nitrite and oxygen

$$NaNO_3 \rightarrow NaNO_2 + O_2$$

$$KNO_3 \rightarrow KNO_2 + O_2$$

- Ammonium nitrate decomposes on heating to give <u>dinitrogen oxide</u> and <u>water</u>

$$NH_4NO_3 \rightarrow N_2O + H_2O$$

- While the nitrite gives nitrogen and water

$$NH_4NO_2 \rightarrow N_2 + H_2O$$

- The metal nitrates (i.e. those of Ca, Mg, Al, Zn, Fe, Pb, Cu) decompose on heating to give the corresponding oxide, nitrogen dioxide and oxygen

$Ca(NO3)_{2(s)} \rightarrow CaO_{(s)} + NO_{2(g)} + O_{2(g)}$

$Mg(NO_3)_{2(s)} \rightarrow MgO_{(s)} + NO_{2(g)} + O_{2(g)}$

- Nitrates of heavy metals (Ag and Hg) decompose to give a metal, $NO_2$ and $0_2$ on heating.

$AgNO_3 \rightarrow Ag + NO_2 + O_2$

$Hg(NO_3)_2 \rightarrow Hg + 2NO_2 + O_2$

### V) Hydroxides

The hydroxide of Na and K are stable to heat ie. Don't decompose on heat even at very high temperature.

All other metal hydroxides decompose on heating to give the corresponding oxides and water vapor

$KOH \xrightarrow[\triangle]{} \text{No Reaction}$

$Mg(OH)_2 \xrightarrow[\triangle]{} MgO + H_2O$

## 5. SOLUBILITY AND SOLUBILITY CURVES

Solubility of a salt in a liquid is the maximum amount of the salt that will dissolve in $100cm^3$ of a liquid at any given temperature

Solubility Curve

These are the graphs which show the variation of solubility with temperature

Solubility of a salt increases with the increase in temperature.

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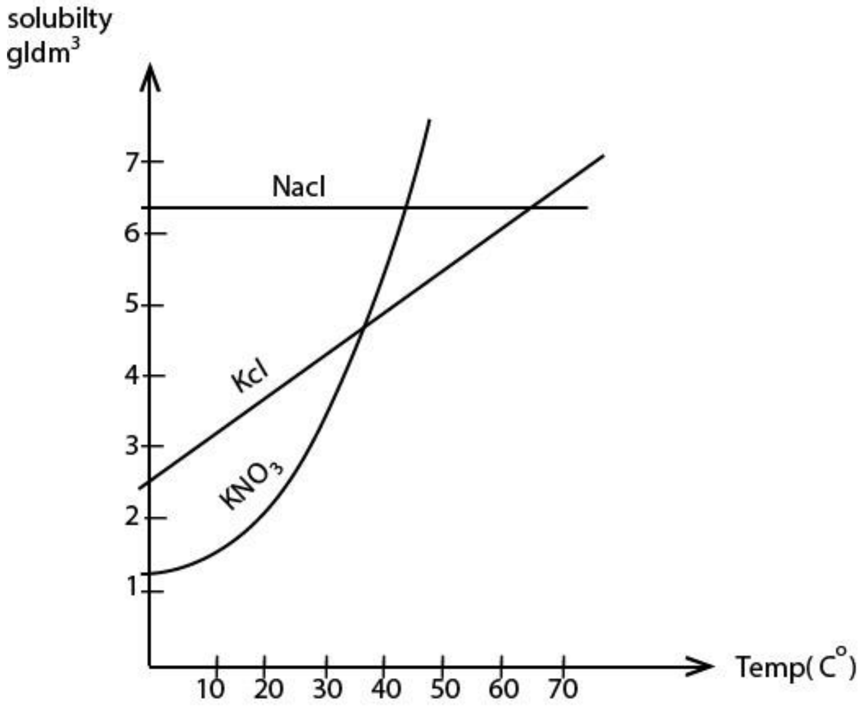
The steeper the solubility curve the more soluble the salt and the easier it is to crystallize that salt

## SOLUBILITY AND SOLUBILITY CURVES.



The diagram above is a graph of solubility against temperature. The vertical component is a solubility of substance in gram per dm$^3$ , where by horizontal component is temperature of a substance. The graph drawn in the plane is analysis of two related variables (solubility and temperature). The change of solubility related to change of temperature.

From the accumulated data due to the test of salts NaCl, KCl and $KNO_3$ analyzed in the graph as drawn above.

(i) A graph of NaCl interpret that, it has a constant solubility at any temperature. The increase or decrease of temperature does not affect its solubility.

(ii) A graph of KCl is a smooth linear graph which interpret that for every change of temperature give the effect change of solubility. Therefore KCl is more soluble than NaCl.

(iii) A graph of $KNO_3$ is a curved graph interpret that the change of temperature gives a rapid solubility. Therefore a salt $KNO_3$ is more soluble than NaCl and KCl.

**Assignment**;

Take a salt of $NaNO_{3(S)}$ to examine its solubility.

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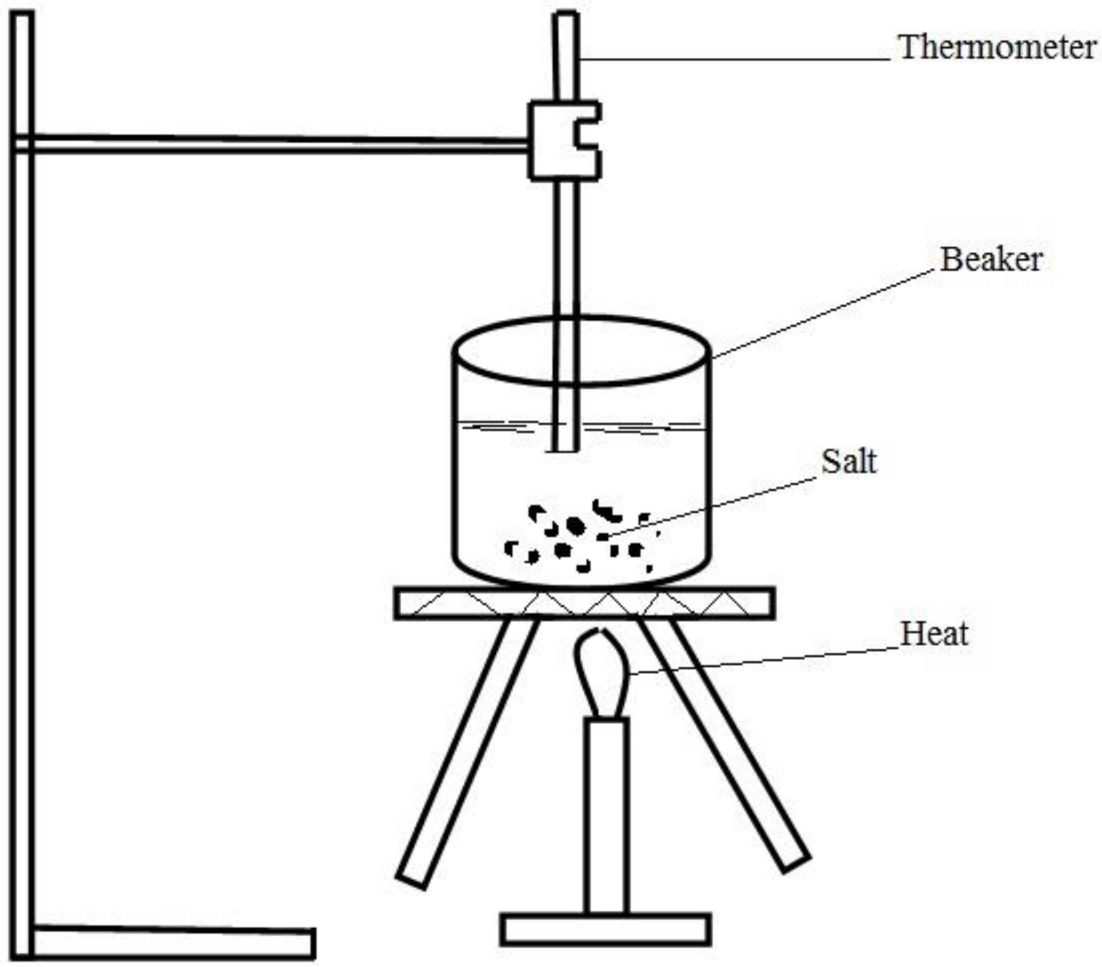
**Procedure**

(i) Measure 1litre of distilled water and pour it into the beaker.

(ii) Measure 1g of $NaNO_3$ by electronic balance and pour into the beaker.

(iii) By heat source, tripod stand, wire gauze, retort stand and thermometer make connection and heat both water and sample salt.



(iv) Watch a disappearance of salt particles until all finish and record the temperature as follows;

| Salt (g) | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 |
|---|---|---|---|---|---|---|---|---|
| Temperature (°C) | | | | | | | | |

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(v) Repeat procedure (ii) for except 2,3,4,5,6,7 and 8

(vi) Draw a graph of solubility against temperature.

## USES OF SALT IN DAILY LIFE

Salt is essential for life, because it has more than 14000 users in daily. But the common uses categorized into Food, Agriculture,Water conditioning, High way Deicing and Industrial chemicals.

### 1. FOOD

i) A salt of sodium chloride is mixed with food as flavoring, (common salt).

ii) A salt of sodium chloride (NaCl)is mixed in the food industrial product as both flavoring and preservative.

iii) A salt of sodium bicarbonate ($NaHCO_3$) 'bicarb' is used in cooking as raising agent for cakes, bread e.t.c.

#### Baking powder

Is a mixture of sodium hydrogen carbonate and tartaric acid, The mixture helps to keep the PH neutral.

### 2. AGRICULTURE

A salt is very important in agriculture since used as land additive nutrients. A fertilizers of Ammonium sulphate ($NH_4SO_{4(S)}$) and Sodium nitrate ($NaNO_3$) salt are used to make land fertility to help healthier growing of crops.

### 3. WATER CONDITIONING

i) In the purification of urban water (Permutit process),the salt of Aluminum silicate is used to remove permanent hardness of water ($Al(SiO_3)_3$).

ii) Sodium carbonate ($Na_2CO_{3(S)}$) called washing soda, is used to make water soft by replacing calcium ions by sodium ions.

iii) Sodium chloride is used in a water softener to generate the ion exchange column.

### 4. HIGH WAY DEICING

A salt of sodium chloride (NaCl) is mixed with grit and spread on roads to prevent road freezing in cold condition.

## 5. INDUSTRIAL USES

i) The salt of potassium Iodide (KI) is added to sodium chloride (common salt) to prevent a lack of Iodine in the diet.

ii) The salt of sodium carbonate is used in the manufacture of glass.

---

## MOLE CONCEPT

### What is mole?

**Mole** is a unit of measurements as any other units,

Example Pair, dozen, gross etc

1mole = $6.02 \times 10^{23}$ particles

**Definition;**

A **mole** is the amount of a substance as many as particles of elementary entities contained in 12g of carbon-12 isotopes. The particles can be atoms, molecules, electrons or ions. This is a unit mole which was introduced by **Avogadro** as **$6.02 \times 10^{23}$ particles**

Example; 1 mole of water ($H_2O$) contains $6.02 \times 10^{23}$ molecules

1 mole of sodium (Na) contains $6.02 \times 10^{23}$ atoms

1 mole of $CuCl_2$ contains $6.02 \times 10^{23}$ ions

1 mole of Fe contains $6.02 \times 10^{23}$ electrons

## MOLAR MASS OF A SUBSTANCE

**Molar mass:** is the mass of I mole of any element or compound

Its SI unit is g/mol. It is denoted by M

For example; Na = its molar mass = 23g/mol

NaCl = its molar mass = 58.5g /mole

$Na_2SO_4$ = its molar mass = (23 x 2) + (32) + (16 x 4) = 142g/mol

NB: Molar masses and relative molecular mass of a substance are calculated from the formula of that substance.

**QUESTION;**

Calculate the molar of the following

1. $Na_2CO_3$. 10 $H_2O$

11. $Al_2 (SO_4)_3$

Al = 27, S = 32, O = 16, Na = 23, C = 12, H = 1

2. Calculate the number of molecules in i. 8g of oxygen gas

ii. 11g of $CO_2$

3. If the number of ions in $CuCl_2$ is 3.02 x 10 $^{23}$, what was the mass of $CuCl_2$

**ANSWERS**:

1. $Na_2CO_3.10H_2O$

(23 x 2) + 12 + (16 x 3) + 10 [(1 x 2) + 16]

46 + 12 + 48 + 180

106 + 180 = 286g/mol

ii) $Al_2 (SO_4)_3$

(27 x 2) + 3 (32 + (16 x 4))

54 + 3(96)

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54 + 288 = 342g/mol

Iii). 8g of oxygen gas

$O_2 = 16 \times 2 = 32 = 6.02 \times 10^{23}$

$8 = ?$

$$\frac{6.02 \times 10^{23} \times 8}{32}$$

$= 1.505 \times 10^{23}$ molecules

IV). 11g of CO2

$CO_2 = 12 + 32 = 44 = 6.02 \times 10^{23}$

$11 = ?$

$$\frac{6.02 \times 10^{23} \times 11}{44}$$

$= 1.505 \times 10^{23}$ molecules

$CuCl_2$

$64 + (35.5 \times 2) = 64 + 71 = 135 = 6.02 \times 10^{23}$

$? = 10^{23}$

$$\frac{3.02 \times 10^{23} \times 135}{6.02 \times 10^{23}}$$

= 67.5 g/mol

## RELATIVE MOLAR MASS ($M_r$)

This is the mass of one molecule of a compound compared with the mass of one atom of carbon – 12.

It has no units

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Molar mass (M) = Relative molecular mass ($M_r$)

Molar mass of an element = relative atomic mass ($A_r$)

Example; Na    M of Na = 23gmol$^{-1}$

$A_r$ of Na = 23

M or $CH_4$ = 16gmol$^{-1}$

Mr of $CH_4$ = 16

M of $H_2O$ = 18gmole

Mr of $H_2O$ = 18

**Example;**

1. Calculating the molar mass of $(NH_4)_2CO_3$ given that N = 14, H= 1, C = 12, O = 16.

Solution

$(NH_4)_2CO_3$

(14g x 2) + (1g x 8) + 12g + (16g x 3)

28g/mol + 8g/mol + 12g/mole + 48g/mol

= 96g/mole

**Molar mass of a compound**

This is the sum of constituent atoms (elements)

**AMOUNT OF A SUBSTANCE OR NUMBER OF MOLES (n)**

Number of moles, n is the mass of the sample of a substance divided by the molar mass of that substance

$N = \frac{m}{M}$ whereby m = mass of a sample

M = Molar mass of substance

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N = number of moles (amount of substance)

**Example;**

What is the amount of substance in

a) 180g of carbon

m = 180g

M = 12gmol$^{-1}$

$$N = \frac{180\ \text{g}}{12\ \text{g/mol}}$$

n = 15mol

b) 180g of $CO_2$

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

$$N = \frac{180\ \text{g}}{44\ \text{g/mol}}$$

n = 4.09 mol

**QUESTIONS:**

Finding the mass of each of the following substance

a) 2.4 mol of NaOH

b) $3.2 \times 10^{-3}$ mol of N

c) 0.780 mol of $Ca(CN)_2$

d) 7.00 mol of $H_2O_2$

How many moles are in each of the following?

a) 0.800g of Ca

b) 79.3g of $Cl_2$

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c) 5.96g of KOH

d) 937g of $Ca(C_2H_3O_2)_2$

## NUMBER OF PARTICLES (N) IN A GIVEN AMOUNT OF SUBSTANCE (n)

To find the number of particles in a given amount of substance we use the expression

N = n.L

Where: n = number of moles

N = number of particles in a substance

L = Avogadro's no = 6.02 x 1023 $mol^{-1}$

**QUESTION:**

1. How many particles are there in 20g of Ca?

Solution

Find n in 20g of Ca

$$N = \frac{20\ g}{40\ g/mol}$$

N = 0.5 mol

From the expression

N = nL

= 0.5 mol x 6.02 x $10^{23}$ $mol^{-1}$

= 3.01 x $10^{23}$

2. How many molecules are there in 80g of the NaOH?

## MOLAR VOLUME ($V_m$)

## MOLAR GAS VOLUME

This is the volume occupied by one mole of a gas at standard temperature and pressure and is 22.4 $dm^3$ or 22.4l

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

Standard temp = $0^0c$ (273K)

Standard pressure = 1 atm (760 mm Hg)

**Avogadro's law** states "At the same temperature and pressure volumes of all gases contain the same number of particles"

In calculating the amount of substance n, using molar volume ($v_m$), the expression used is

$$\text{Amount of substance n} = \frac{\text{volume of gas}(dm^3\ )}{\text{Molar volume}(\frac{dm^3}{mol})}$$

$$\text{Amount of substance} = \frac{v}{22.4\frac{dm^3}{mol}}$$

Example; Find the amount of substance present in

a) 18.8 $dm^3$ of $CO_2$ at s.t.p

b) 48.8 $dm^3$ of $O_2$ at s.t.p

Solution

$$\text{From the expression } n = \frac{v}{v_m}$$

Where; V = volume of a gas

$V_m$ = molar volume of a gas

n = amount of substance

V = 18. 8 dm3

$V_m$ = 22.4 dm3/ mol

n =?

Substituting the volume in the expression

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$$\frac{18.8}{22.4}$$

n = 0.839 mol

b) $N = \frac{48.8dm^3}{22.4dm^3/mol}$

n = 2.18 mol

## THE MOLE IN STOICHIOMETRIC CALCULATIONS

A balance chemical equation tells us a great of quantitative information

Consider the following equation us a great deal of quantitative information

Consider the following balanced equation

$$2H_6 + O_2 \rightarrow 2H_2O$$

$$4g + (16x2 = 332g) \rightarrow 36g$$

2moles 1mol 2mol

Stoichiometry- is the quantitative relationship of reacting substances

Stoichiometric coefficients = moles

A balanced equation is used in calculating different quantities

Example; $CaCO_3 \rightarrow CaO + CO_2$

From the above balanced equation, calculate the weight of CaO and Volume of $CO_2$ at s.t.p which will be produced by heating 75g of $CaCO_3$

(Ca = 40, C = 12, O = 16)

Solution

Finding the molecule mass

$CaCO_3 \rightarrow CaO + CO_2$

$(40 + 12 + 48)\ (40 + 16)$

Molecular mass of $CaCO_3$ = 100g

CaO = 56g

From the equation:

$CaCO_3 \rightarrow CaO + CO_2$

1 mol 1mol 1mol

100g 56g

15g x?

$100\text{g produces } 56\text{ g} = \frac{100\text{ gx}}{100\text{g}} = \frac{56\text{g x } 75\text{g}}{100\text{ g}}$

75g produces ?

x=42g

If 100g/mol = 22.4dm$^3$

75g/mol =?

= 16.8 dm$^3$

**QUESTION**

a) How many ions are in 10g of $Al_2(SO_4)_3$?

b) How many fluorides are in 1.46mols of $AIF_3$?

**Solution;**

1(a) $Al_2(SO_4)_3 \rightarrow 2Al^{3+} + 3SO_4^{2-}$

342g/mol 2mol 3mol

342 g/mol = 5 mol of ions

10 g = ?

If 342 g = 1 mol

10 g = x

x = 0.0292 mol of $Al_2 (SO_4)_3$

If 1 mol of $Al_2 (SO_4)_3$ = 5mol of ions

0.0292mol of $Al_2 (SO_4)_3$ = ?

$$\frac{0.0292 \text{ mol x } 5 \text{ mol of ions}}{1 \text{ mol}}$$

= 0.0146 mol of ions

N = nL

= 0.046 mol x 6.02 x 1023 mol-1

= 8.79 x $10^{23}$

## SOLUTION AND MOLAR CONCENTRATION

Solution is a mixture of solvent and solute.

**Solute** - Is a solid crystal component which dissolved into solvent.

**Solvent** - Is a liquid component which dissolve solute. (water is a common universal solvent).

So that when we prepare a standard solution, we must measure both two component. A solute measured in grams of weight and solvent in liter, dm3 or cm3 of its volume. The quantity of solute that may dissolved by 1litre of solvent make a standard solution.

### STANDARD SOLUTION

Is a solution of known concentration. Standard solution has been prepared by accurate measurement and the concentration of solute in the solution is described into two ways.

i. Weight concentration in grams.

ii. Particles concentration in mole.

### i. WEIGHT CONCENTRATION

This concentration abbreviated as conc. of a grams dissolved in one litre (g/dm3 or g/L).

e.g 30g/L or 15g/dm3

**Example;**

Allen dissolved two spoons of sugar into the tea drink. If capacity of cup is 200cm3 and a spoon carry 25g, what is the concentration of sugar in the tea?

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**Solution**

carrying capacity = 25g

25 x 2 = 50g

Therefore, weight of solute = 50g

Volume of solvent = 200cm3

But,

$$50\text{g} \longrightarrow 200\text{cm}^3$$

$$\text{x?} \longrightarrow 1000\text{cm}^3$$

$$\text{x} = \frac{50\text{g} \times 1000\cancel{\text{cm}^3}}{200\cancel{\text{cm}^3}} = 250\text{g}$$

$$\text{Therefore, conc.} = 250\text{g/L}$$

**MOLARITY (MOLAR CONCENTRATION)**

This is a concentration of particles contained in one litre of the solution. (mol/L or M)

**Example;**

The Allen spin a common salt in the bowl filled by 320cm3 of water to make a solution. If the same spoon used, determine a molarity of salt in the solution.

**Solution**

Volume of water = 320cm3

Weigh of salt = 25g

But, Moles = weight of component/Molar mass (W/Mr)

n = 25g/58.5g

n = 0.43moles

Therefore,

$$0.43\text{moles} \longrightarrow 320\text{cm}^3$$

$$\text{x?} \longleftarrow 1000\text{cm}^3$$

$$\text{x} = \frac{0.43\text{moles} \times 1000\cancel{\text{cm}^3}}{320\cancel{\text{cm}^3}} = 1.3$$

$$\text{Therefore, Molarity} = 1.3\text{mol/L}$$

**RELATIONSHIP OF WEIGHT AND MOLAR CONCENTRATION**

From,

Molarity = moles/Volume in litre

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M = n/V in L

But,

n = W/Mr

Therefore,

M = W/Mr X 1/V in L

But,

W/V in L = Conc

Therefore,

M = Conc/Mr

So that,

Molarity = Conc/ Molar mas

Where by,

n = no of moles

V = Volume in liter

Mr = Molar mass

W = Weight of component in grams

**Assignment**

Find the molarity of each component if both had a concentration of 72g/dm3 .

(a) $H_2SO_4$
(b) $Na_2CO_3$
(c) $NaHCO_3$
(d) KCl

---

## QUANTITATIVE ANALYSIS AND VOLUMETRIC ANALYSIS

### Definition;

**Volumetric analysis,** is the process of determining the amount of substance in term of volume and concentration of reacting solutions.

### Concentration;

This is the amount of substance per one $dm^3$ (litre) of solution

Units are $g/dm^3$, $mole/dm^3$

### Types of concentration

1. **Molar concentration (morality)** is the amount of substance in moles per one $dm^3$ of solution ($mol\ dm^{-3}$) it is denoted by M

For example; 0.1M $H_2SO_4$, 2M $H_2SO_4$ etc

Mass concentration is the amount of substance in games per one $dm^3$ of solution ($g/dm^3$)

**Question:**

a) Define (I) concentration (ii) molarity

b) Calculate mass concentration of the following

i) 4g of NaOH in $400cm^3$ of solution

ii) 10g of $H_2SO_4$ in $2dm^3$

iii) 2g of HCl in 4l

Solution

$$\text{Given } m = 4g,\ \text{Volume} = \frac{400cm^3}{1000}$$

$$= 0.4dm^3$$

$$\text{Mass concentration} = \frac{4g}{0.4dm^3}$$

$$= 10gdm^{-3}$$

a) Give the meaning and SI units of molarity

b) What is the molarity of a solution containing?

i) 0.05mole in 80 $cm^3$

ii) 0.25mole in $2dm^3$

iii) 0.5mole in 4l

**Solution**

2b). given n = 0.05mol

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$V = \frac{80}{1000} = 0.08\text{dm}^3$

$M = \frac{0.05}{0.08}$

$= 0.625\text{mol/dm}^3$

## STANDARD VOLUMETRIC APPARATUS

The apparatus needed in volumetric analysis include

- Retort stand and clamps
- White tiles
- Burette – This is calibrated from 0 to 50 $\text{cm}^3$ it measures volume of an acid
- Pipette – common capacity volumes are 25 $\text{cm}^3$ or $20\text{cm}^3$. It measures volume of a base.
- Conical flask
- Filter funnel
- Reagent bottles
- Volumetric flask
- Wash bottle
- Measuring cylinder
- Beakers
- Watch glass

## STANDARD SOLUTION

A standard solution is the one whose concentration is exactly known

For example; 10g of NaOH in $1\text{dm}^3$ of solution

10g of KOH in $2\text{dm}^3$ of solution

The molarity of the above substance can be calculated because their concentration (in $\text{gdm}^{-3}$) and formulae are known

$$\text{Molarity} = \frac{\text{concentration}\left(\frac{\text{g}}{\text{dm}^3}\right)}{\text{molar mass}\left(\frac{\text{g}}{\text{mol}}\right)}$$

## MOLAR SOLUTION

A molar solution is the solution which contains one mole of a solute dissolve in one $dm^3$ of solution

Example; 40g of NaOH dissolved in $1dm^3$ of solution

98g of $H_2SO_4$ dissolved in 1 litre of solution

132g of $(NH_4)_2SO_4$ dissolved in 1L of solution

## TITRATION OF ACIDS AND BASES

**Titration**: is the process of adding an acid to a base until an indicator shows that the reaction is complete

What is acid base titration?

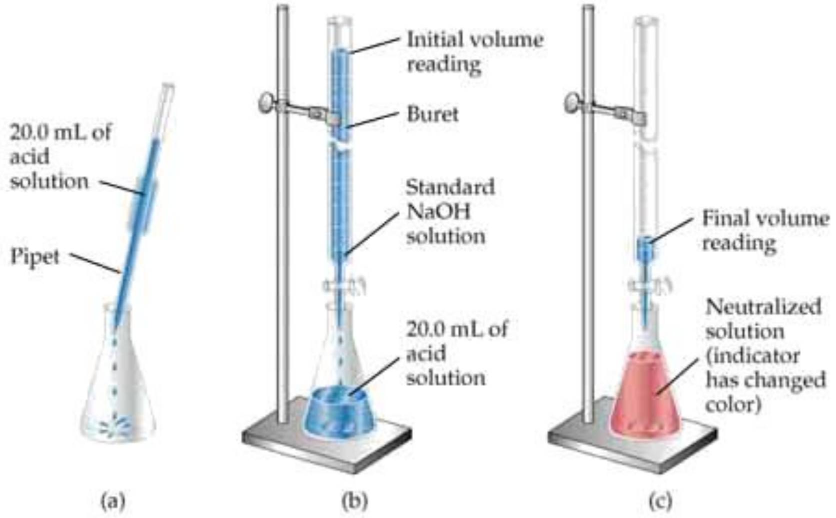
What is titration?

This is a process of adding one solution to another until an indicator shows that the reaction is complete

## CHOICE OF INDICATORS IN ACID – BASE TITRATION

| ACID/ BASE | EXAMPLE | SUITABLE INDICATOR |
|---|---|---|
| Strong acid/ strong base | HCl / NaOH | Any indicator |
| Weak acid/ strong | Acetic acid/$CH_3COOH$/ KOH | Phenolphthalein p.o.p |
| Strong acid/ weak base | $H_2SO_4$/$NH_3$ /($NH_4OH$) | Methyl orange |
| Weak acid/ weak base | $CH_3COOH$/$NH_4OH$ | No satisfactory indicator |

In an acid – base titration, the solution of the known concentration is added gradually to a solution of unknown concentration. When the unknown solution is exactly neutralized as shown by the color change of the indicator then at this point the number of moles of the solution with known concentration, for example; acid is equal to the number of moles of solution with unknown concentration, for example; Base. This point is known as End – point of **titration**.



## MOLAR RATIO

The amount of moles present in the two different solution A and B can be symbolized as $n_A$ and $n_B$.

$$\text{The ratio of moles} = \frac{nA}{nB}$$

But the moles in standard solution,

$n = M \times V$

Where by,

M = Molarity

V = Volume taken

Therefore,

$n_A = M_A \times V_A$

$n_B = M_B \times V_B$

So that,

$$\frac{M_A V_A}{M_B V_B} = \frac{n_A}{n_B}$$

This is called molar ratio formula which can be used in titration to determine any unknown, if five items are known. But always $n_A$ and $n_B$ identified from the balanced chemical equation of compound A and B.

**QUESTION**

1. 24cm$^3$ of a solution of 0.1M KOH were exactly neutralized by 30cm$^3$ of solution of sulphuric acid

i) mol/dm$^3$ ii) g/dm$^3$

**Solution;**

Make the equation;

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

$n_B = 2$ $n_A = 1$

$V_B = 24cm3$ $V_A = 30cm3$

$M_B = 0.1M$ $M_A = ?$

$$M_A = \frac{M_B V_B n_A}{V_A n_B}$$

$M_A = 0.04mol/dm^3$

ii) Concentration of acid in $g/dm^3$

$$\text{Molarity} = \frac{\text{mass concentration}}{\text{molar mass}}$$

Mass concentration = molarity x molar mass

= 0.04mol/dm3 x 98g/mol

= 3.92 $g/dm^3$

**ASSIGNMENT:**

**LAMBAT GENERAL QUESTION 29 ON WARDS**

**QUESTION**

$25cm^3$ of $X_2CO_3$ contains $10.6g/dm^3$ required 23.00 $cm^3$ and 0.217M of HCL for complete neutralization calculating

i)The molarity of $X_2CO_3$ solution

ii) The molar mass of $X_2CO_3$

iii) The atomic mass of X

iv)Name of element X

**Solution**

i) $X_2Co_3 + 2HCL \longrightarrow 2xCl + H_2O + CO_2$

$n_B = 1$ $M_A = 0.217$

$V_B = 25cm3$ $n_A = 2$

$M_B = ?$ $V_A = 23cm^3$

$n_A = n_B$

$M_A V_A = M_B V_B$

$$\frac{M_B V_B n_A}{V_B n_A} = \frac{M_A V_A n_B}{V_B n_A}$$

$$M_B = \frac{M_A V_A n_B}{V_B n_A}$$

$$M_B = \frac{1 \text{ x } 0.217 \text{ x } 23}{2 \text{ x } 25} = \frac{0.217 \text{ x } 23}{50}$$

$= 0.09982mol/dm^3$

ii) $\text{Molarity} = \dfrac{\text{mass concentration}}{\text{molar mass}}$

0.09982 = 10.6

1 = M

$$= \frac{10.6 \text{ x } 100000}{0.09982 \text{ x } 100000}$$

= 106g/mol = 106.2g/mol

iii) $X_2\ CO_3$

2x + 12 + 48 = 106

$2x + 60 = 106$

$2x = 46$

$x = 23$

$= 23g$

iv)The name of element X is Sodium (Na)

## PERCENTAGE COMPOSITIONS OF PURE AND IMPURE COMPONENTS

The pure substance may be composed in pure substances without affecting its chemical properties. Through volumetric analysis a percentage composition may be determined.

- Percentage purity
- Percentage impurity.

Where by,

$$\% \text{ Purity} = \frac{\text{Pure Conc.}}{\text{Impure Conc.}} \times 100\%$$

Density = mass concentration per $cm^3$

**EXAMPLES;**

3g of impure $Na_2CO_3$ were made up to 250 $cm^3$ of solution; 25$cm^3$ of this solution required 21.0 $cm^3$ of 0.105M HCl for complete neutralization

i) Write down the balanced equation for the reaction

ii) Calculate the molarity of pure $Na_2CO_3$

iii) Calculate concentration of pure $Na_2CO_3$

iv) Calculate % purity of $Na_2CO_3$

v) Calculate % impurity of $Na_2CO_3$

Solution i and ii

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 + H_2O$$

$n_B = 1$ $n_A = 2$

$M_B = ?$ $M_A = 0.105$

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$V_B = 25cm^3$ $V_A = 21cm^3$

$n_A = n_B$

$M_AV_A = M_BV_B$

$$M_B = \frac{0.105 \times 21 \times 1}{25 \times 2}$$

= 0.0441 M (pure $Na_2CO_3$)

It's pure because sand wouldn't react with the acid

(iii) Concentration = molarity x molar mass

= 0.0441 M x 106g/mol

= 4.67g/$dm^3$

iv) 3.0g=250$cm^3$

x = 1000$cm^3$

$$x = \frac{3 \times 1000}{250}$$

x = 12 g/$dm^3$

$$\% \text{ Purity} = \frac{\text{concentration of pure}}{\text{concentration of impure}} \times 100\%$$

$$\text{Percentage purity} = \frac{4.67}{12} \times 100\%$$

Percentage purity = 38.9%

Percentage impurity= 100% - 38.9%

= 61.1 %

DILUTION LAW.

Is the law that describe or direct how to change and decrease molar concentration of the solution. The law state that “The product of initial concentration and its volume is equal to the product of final concentration and its volume”

$M_1 V_1 = M_2V_2$

Where, $M_1$ = initial concentration

$M_2$ = final concentration

$V_1$ = initial volume

$V_2$ = final volume

NB: $M = \frac{n}{v}$

$M \propto \frac{1}{v}$

**Example.**

1. To what volume must 300cm$^3$ of 6.0M NaOH be diluted to give a 0.40M solution?

**Solution**

$M_1 = 6.0M$ $V_1 = 300cm^3$

$M_2 = 0.04M$ $V_2 =?$

$M_1V_1 = M_2V_2$

$V_2 = \frac{M_1V_1}{M_2}$

$V_2 = \frac{6 X 300}{0.4}$

$\underline{V_2 = 4500cm^3}$

**DETERMINING WATER OF CRYSTALLIZATION.**

Water of crystallization is the water that is bound within the crystals of substances. This water can be removed by heating. titration reaction are used to determine the amount of water of crystallization in a substance.

**Questions**

1. Determine the number of molecules of water of crystallization in the sample ($Na_2CO_3 . XH_2O$). If R.M.M is 286g/mol.

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**Solution**

$Na_2CO_3 . XH_2O$
$2(Na) + C + 3(O) + X(H_2O) = 286$
$2(23) + 12 + 3(16) + X(18) = 286$
$106 + X(18) = 286$
$18X = 286 - 106$
$18X = 180$
$X = 180/18$
$X = 10$

Therefore, the number of moles of water = 10

2. a) How many grams of concentrated HBr should be used to prepare 500ml of 0.6 M HBr solution?

b) Find the volume of stored solution.

**Solution**

a) Given purity = 48%

D = 1.5g

V = 500ml

M = 0.6M

n = M x V

= 0.5l x 0.6mol/l

= 0.3mol

$$N = \frac{\text{mass}}{\text{molar mass of HBr}}$$

m = 0.3mol x 81g/mol = 24.3g

$$\%\text{purity} = \frac{\text{maspure x 100}}{\text{Impure}}$$

$$48 = \frac{24.3\text{g x }100}{x}$$

$$x = \frac{24.3 \text{ x } 100}{48}$$

x = 50.6g

b) D = 1.5g/ml

m = 50.6g

v =?

$$V = \frac{50.6g \times 10}{41.5g/ml \times 10}$$

V = 33.7ml

3. 1.5g of an acid Hx was dissolved in water and its solution made to 250cm$^3$. If 30.2cm3 of this acid solution neutralized 25cm3 of 0.115M KOH, Calculate,

i) Molarity of Hx

ii) Molecule weight of Hx

iii) Name of radical x

**Solution**

i) $V_A$ = 30.2cm$^3$ $V_B$ = 25cm$^3$

$M_A$ =? $M_B$ = 0.115M

Mass of HX = 1.5g

KOH + HX KX + $H_2O$

$n_B$ = 1

$n_A$ = 1 thus $M_A$ $V_A$ = $^{MB}$

$$M_A = \frac{M_B V_B}{V_A} = \frac{0.115M \times 25cm^3}{30.2cm^3}$$

=

0.095M

ii) $\frac{\text{Mass concentration}}{\text{Molar weight}} = \text{Molarity}$

$1.5g \rightarrow 2.50cm^3$

$? \rightarrow 1000cm^3$

$X = 6g/dm^3 \times 1000$

$\text{Molarity} = \frac{\text{mass concentration}}{\text{molar mass}}$

Molar weight =
63g/mol

iii) x must be a radical and the common ones are HCl, $HNO_3$, $H_2SO_4$, $H_3PO_4$, so it's either $HNO_3$ or HCl.

But $HNO_3$ = 63g/mol

Thus, x is nitrate (v) ion ($NO_3^-$)

---

## IONIC THEORY AND ELECTROLYSIS

### ELECTROLYSIS.

Definition: is the decomposition of electrolytes by the passage of an electric current through it.

**Electrolyte**:

This is the substance which either in solution or in molten state conducts electric current to undergo decomposition due to its passage.

**For example;** $H_2SO_4$ - Sulphuric acid

$HNO_3$ – Nitric acid

NaCl – Sodium chloride

Salts, acids and bases

**A non – electrolyte**

This is a substance which either in solution or molten state doesn't conduct an electric current

Example Sugar, ethanol, urea etc.

Types of electrolytes

There are two types of electrolytes

(a) **Strong electrolyte**

Examples of Strong bases; $NaOH \rightarrow Na^+ + OH^-$

Strong acids; $HCl \rightarrow H^+ + Cl^-$

Salts; $Na_2CO_3 \rightarrow Na^+ + CO_3^{2-}$

These are electrolytes which ionize completely in solution example $H_2SO_4$, HCl, NaCl, and NaOH

b) **Weak electrolytes**

These are electrolytes which ionize partially in solution. Examples are,

Weak acids such as $CH_3COOH$, $H_2CO_3$

Water

Weak bases such $NH_3$ solution or $NH_4OH$

**MECHANISM OF ELECTROLYSIS**

When an electrolyte is dissolved in water it dissociates to form ions thus the aqueous solution of electrolyte contains positive and negative ions

Example strong electrolyte

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$$NaCl_{(aq)} \longrightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

$$H_2SO_{4(aq)} \longrightarrow 2H^{+}_{(aq)} + SO_4^{2-}{}_{(aq)}$$

$$H_3PO_{4(aq)} \longrightarrow 3H^{+}_{(aq)} + PO_4^{3-}{}_{(aq)}$$

Weak electrolytes

$$\left.\begin{array}{l} CH_3COOH \rightleftharpoons CH_3COO^- + H^+ \\ NH_4OH \rightleftharpoons NH_4^- + OH^- \\ H_2O \rightleftharpoons H^+ + OH^- \end{array}\right\} \text{Partial ionization}$$

**IONS:**

Ion is a charged particle

It’s either positively charged or negatively charged

a) **Anions** - are negatively charged ions

b) **Cations** – are positively charged ions

## THE ELECTRODES:

These are the two poles of metals rods or carbon by which the electrons leave or enter the electrolyte

**Types of electrodes**

a) **The Anode** (+)

This is the positive electrode by which the electrons leave the electrolyte

b)**The cathode** (-)

This is the negative electrode by which electrons enter the electrolyte

When an electric current passes through an electrolyte, the ions move to the oppositely charged electrodes.

The Cations (positive ions) move to cathode

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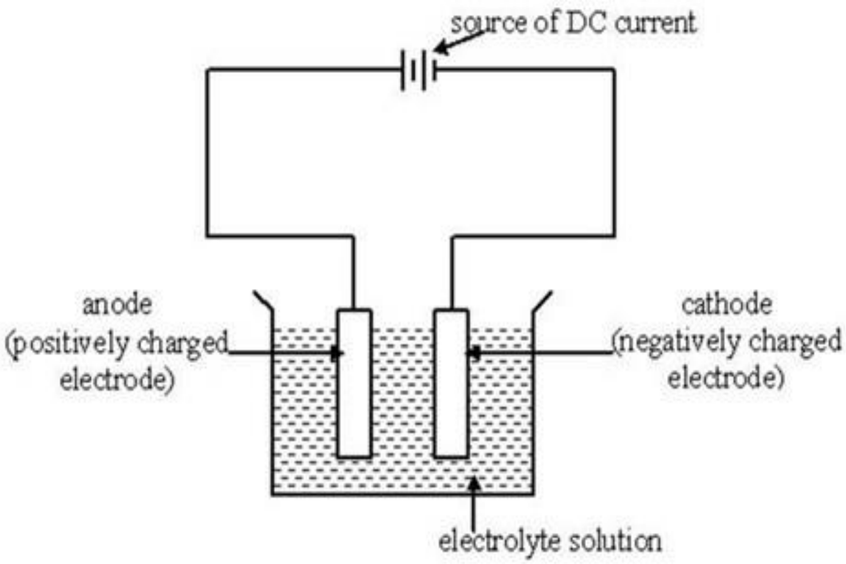
The anions (negative ions) move to anode



## ELECTRODE REACTIONS

### 1. Anode reaction

Anions carry negative charges due to extra electrons they have.

On arrival at the surface of anode they surrender the extra electron. By so doing they become discharged.

I. Discharged of univalent anions; $X^- \rightarrow X + e^-$

Ii Discharged of divalent anions; $X^{2-} \rightarrow X + 2e^-$

Iii Discharge of trivalent anions; $X^{2-} \rightarrow X + 2e^-$

Iv Discharge of tetravalent anions; $X^{4-} \rightarrow X + 4e^-$

Any process involving the loss of electrons is called oxidation. In electrolysis the reaction at the anode is called oxidation.

Therefore,another definition of oxidation;

This is the process whereby the substance looses electrons.

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**2. Cathode reaction**

Cations have positive charges because they have deficiency of electrons. On arrival at the surface of cathode, they receive electrons from the cathode.

By so doing they become discharged

Example; i) Univalent cations discharged as

$$M^+ + e^- \rightarrow M$$

ii) Divalent Cations discharge as

$$M^{2+} + 2e^- \rightarrow M$$

iii) Trivalent Cations discharge as

$$M^{3+} + 3e^- \rightarrow M$$

Any process involving the gain of electrons is called reduction. The cathode reaction is reduction process.

Therefore, reduction is the process whereby the substance gains electrons.

## PREFERENTIAL DISCHARGE OF IONS:

When two or more ions of similar charge are present under similar conditions in a solution.

2g $H^+$ and $Na^+$ or $OH^-$ and $SO_4^{2-}$, one of them is preferentially selected for discharge

The selection of ion to discharge depends on the following factors;

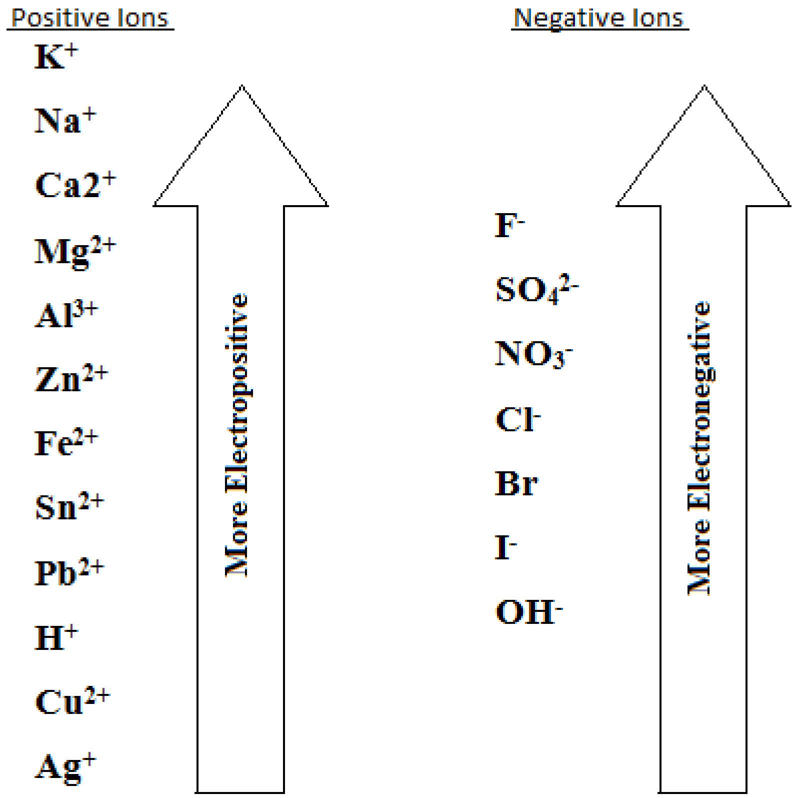
a) Position of the metal or radical in the electrochemical series

b) Concentration

c) Nature of the electrode

## A POSITION OF THE METAL/GROUP IN ELECTROCHEMICAL SERIES

The electrochemical series is an arrangement of metal ions in the order which can be reduced (discharged). Those which can be reduced (discharged) with difficult are placed at the top of the series example $K^+$, $Ca^{2+}$ and $Na^+$ may accept electrons with great difficulty. $Ag^+$ and $Cu^{2+}$ can accept electrons readily/ easily.



The ion tower in electrochemical series is discharged in preference, to those above it

**Example;** electrolysis of dilute hydrochloric acid using platinum electrode ions present

$$HCl_{(aq)} \rightarrow H^+ + Cl^-$$

$$H_2O_{(l)} \rightarrow H^+ + OH^-$$

Cathode reaction

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$H^+$ will migrate to the cathode and be discharged

$$H^+ + e^- \rightarrow H$$

$$\underline{H^+ + e^- \rightarrow H}$$

$$2H^+ + 2e^- \rightarrow H_{2(g)}$$

$H_2$ gas will evolve at the cathode

Anode reaction

$Cl^-$ and $OH^-$ will migrate to and $OH^-$ will be preferentially discharged because it is lower than $Cl^-$ in electrochemical series

$$OH^- \rightarrow OH^- + e^-$$

$$\underline{OH^- \rightarrow OH^- + e^- \quad}$$

$$2OH^- \rightarrow H_2O + O + 2e^-$$

$$\underline{2OH^- \rightarrow H_2O + O + 2e^-}$$

$$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$$

$O_2$ gas will evolve at the anode

**Conclusion:** The electrolyte concentration will be increasing as the elements of water are the ones coming out as gases

## ELECTROLYSIS OF DILUTE SULPHURIC ACID (using platinum electrode)

## ELECTOLYSIS OF WATER

Ions present; $H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$

$$H_2O \rightarrow H^+ + OH^-$$

Anode reaction

$SO_4^{2-}$ and $OH^-$ will migrate to the anode and $OH^-$ will be preferentially selected to discharge

$OH^- \rightarrow OH^- + e^-$

$\underline{OH^- \rightarrow OH^- + e^-\quad}$

$2OH^- \rightarrow H_2O + O + 2e^-$

$\underline{2OH^- \rightarrow H_2O + O + 2e^-}$

$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$

$O_{2\,(g)}$ will evolve at the anode

Cathode reaction

$H^+$ will migrate to the cathode and discharge

$H^+ + e^- \rightarrow H$

$\underline{H^+ + e^- \rightarrow H}$

$2H^+ + 2e^- \rightarrow H_{2(g)}$

Hydrogen will evolve at the cathode

**Overall effect**

The concentration of the electrolyte will increase because during electrolysis the elements of water comes out as gases

## CONCENTRATION OF THE IONS

Increase in concentration of an ion(s) tends to promote its discharge

Example Electrolysis of concentrated hydrochloric acid.

Ions present; $HCl \rightarrow H^+ + Cl^-$

$H_2O \rightarrow H^+ + OH^-$

Anode reaction

$Cl^-$ and $OH^-$ will migrate to the anode and $Cl^-$ will be discharged in preference to the $OH^-$ because it's available in higher concentration.

$Cl^- \rightarrow Cl + e^-$

$\underline{Cl^- \rightarrow Cl + e^-}$

$2Cl^- \rightarrow Cl_2 + 2e^-$

Chlorine gas will evolve at the anode

Cathode reaction

$H^+ + e^- \rightarrow H$

$\underline{H^+ + e^- \rightarrow H}$

$2H^+ + 2e^- \rightarrow H_{2(g)}$

Hydrogen gas will evolve at the cathode

Overall effect.

-Concentration of electrolyte will be decreasing because elements of HCl are moving out as gases.

NB: whenever a halogen ($Cl^-$, $Br^-$, $I^-$) is involved in competition for discharge, the concentration factor holds an upper hand especially chloride salts, salts of Halogens e.g. $PbBr_2$

**NATURE OF THE ELECTRODE**

In some reactions of electrolysis the type of electrode used will determine the products of electrolysis.

Consider the following examples

1. a) Electrolysis of sodium chloride by using carbon.

$$\text{Ions present; } NaCl_{2(aq)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

$$H_2O_{(l)} \rightarrow H^{+}_{(aq)} + OH^{-}_{(aq)}$$

At the cathode both $Na^+$ and $H^+$ will migrate to the cathode. $H^+$ will be selected to discharge in the preference to $Na^+$ due to $Na^+$ position in electrochemical series

Cathode reaction

$$H^{+} + e^{-} \rightarrow H$$

$$\underline{H^{+} + e^{-} \rightarrow H}$$

$$2H^{+} + 2e^{-} \rightarrow H_{2(g)}$$

At the anode

Both $Cl^-$ and $OH^-$ will migrate to the anode and $Cl^-$ will be discharged due to the concentration factor

$$Cl^{-} \rightarrow Cl + e^{-}$$

$$\underline{Cl^{-} \rightarrow Cl + e^{-}}$$

$$2Cl^{-} \rightarrow Cl_2 + 2e^{-}$$

$H_{2\,(g)}$ will evolve at the cathode and

$Cl_{2\,(g)}$ will evolve at the anode

B) Electrolysis of sodium chloride by using mercury cathode

$$\text{Ions present; } NaCl_{2(aq)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

$$H_2O_{(l)} \rightarrow H^{+}_{(aq)} + OH^{-}_{(aq)}$$

| **At the cathode** | **At the anode** |
|---|---|

| | |
|---|---|
| Both $H^+$ and $Na^+$ will move to the cathode of mercury. But it's $Na^+$ which will be discharged due to the nature of electrode. | Both $Cl^-$ and $OH^-$ will move here and due to the concentration factor. $Cl^-$ will be discharged in the preference to $OH^-$ |

| Cathode reaction | Anode reaction |
|---|---|
| $Na^+ + e^- \to Na$ | $Cl^- \to Cl + e^-$ |
| Sodium metal will be deposited | $\underline{Cl^- \to Cl + e^-}$ |
| | $2Cl^- \to Cl_2 + 2e^-$ |
| | Chloride gas will involve at the anode |

2(a) Electrolysis of $CuS0_4^{2-}$ by using Carbon / platinum electrode:

Ions present; $CuSO_4 \to Cu^{2+} + SO_4^{2-}$

$$H_2O \to H^+ + OH^-$$

At the cathode

Ions present are $Cu^{2+}$ and $H^+$

At cathode reaction

$$Cu^+ \to 2e^- + Cu$$

Because $Cu^{2+}$ is lower in e.c.s copper metal will be deposited at the cathode

At the anode.

Ions present are $SO_4^{2-}$ and $OH^-$

Anode reaction

$$4OH^+ \to 2H_2O + O_2 + 4e^-$$

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Oxygen gas will be produced at the anode.

**Overall effect**

a) Reddish brown copper metal is deposited on cathode

b) $O_{2\,(g)}$ is liberated at the anode

c) Concentration of $CuSO_4$ decrease i.e. blue colour of $CuSO_4$ turns to colourless since the copper II sulphate is converted to form sulphuric acid or the number of $Cu^{2+}$ in the solution decreases

2(b) Electrolysis of $CuSO_4$ using copper electrodes

$$\text{Ions present; } CuSO_4 \rightarrow Cu^{2+} + SO_4^{2-}$$

$$H_2O \rightarrow H^+ + OH^-$$

At the cathode

Ions moved are $Cu^{2+}$ and $H^+$ cathode reaction

$$Cu^{2+} + 2e^- \rightarrow Cu$$

Copper metal is produced at the cathode

At the anode

Ions moved here are $SO_4^{2-}$ and $OH^-$

Anode reaction

When copper anode is used there are 3 possibilities

I) Discharged of $SO_4^{2-}$

ii) Discharge of $OH^-$

Iii) Conversion of Cu atoms to $Cu^{2+}$

The conversion of Cu to $Cu^{2+}$ occurs most readily

Anode reaction therefore is

$$Cu \rightarrow Cu^{2+} + 2e^{-}$$

Copper ions are released into the solution

Overall effect

a) Reddish brown copper metal is deposited at the cathode

b) Copper ions released into the solution from anode

c) Concentration of $CUSO_4$ remains constant since the number of $Cu^{2+}$ discharged at the cathode are replaced by the released $Cu^{2+}$ from the anode. Hence no colour change.

d) Anode decreases in weight

This electrolysis is only the transfer of copper from anode to cathode.

## LAWS OF ELECTROLYSIS;

A man called faraday put forward two laws expressing results of experiments of electrolysis. These laws they assert that the amount of an element liberated during electrolysis depends on

i) The time of passing the steady current

ii) The magnitude of the steady current

iii) The charge on the ion of the element

## FARADAY'S FIRST LAW OF ELECTROLYSIS

It states that;

"The mass of a substance liberated at or absorbed from an electrode during electrolysis is directly proportional to the quantity of electricity passed through the electrolyte"

If symbols are used the first law can be stated mathematically as

$$M \propto It$$

Where;m = mass of the substance liberated

t = time in seconds

I = current in Amperes.

To replace the symbol of proportionality by an equal sign we must introduce a proportionality constant Ƶ

m = zIt

But It = Q, where Q = quantity of electricity

m = ƵQ

Ƶ = is known as electrochemical equivalent of an element

**Definition;**

**The electrochemical equivalent** of a substance is the mass of a substance that can be liberated by passage of one coulomb

Its SI unit is **g/c**

Every element has its own value of electrochemical equivalent of an element

To calculate the value of Ƶ of an element we use the formula

$$Ƶ = \frac{Ar}{V.F}$$

Where; Ar = relative atomic mass of an element

V = valency of the element

F = faraday's constant (96500c)

Examples;

| Element | Ar | Valency | Z |
|---|---|---|---|
| Cu | 64 | 2 | $\frac{64}{2 \text{ X } 96500} = 3.32 \times 10^{-4}$ g/c |
| Zn | 65 | 2 | $\frac{65}{2 \text{ X } 96500} = 3.38 \times 10^{-4}$ g/c |
| Al | 27 | 3 | $\frac{27}{2 \text{ X } 96500}$ = 9.32 x 10-4 g/c |

NB:Faraday's constant is the quantity of electric charge carried by electrons

(It is 96500 C) per mole

**FARADAY'S SECOND LAW OF ELECTROLYSIS**

It states that

"When the same quantity of electricity is passed through solutions of different electrolytes, the amounts of elements deposited are proportional to the chemical equivalent of the element"

Chemical equivalent of the element is obtained by dividing it's relative atomic mass by the charge of its ion

Example Chemical equivalent

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

$$\text{Example Z of } Ag^{+1} = \frac{108}{1}$$

$$Zn = \frac{65}{2}$$

$$Na = \frac{23}{1}$$

Mathematically

Faraday's law,

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

**MATHEMATICAL STATEMENT OF FARADAY'S 2$^{ND}$ LAW OF ELECTROLYSIS**

$n \propto 1/v$

$m \propto E$

Where v = charge of ion

m = C.E

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

$$\frac{m_1}{E_1} = C \qquad \frac{m_2}{E_2} = C$$

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

$$\frac{m_1}{m_2} = \frac{E_1}{E_2}$$

$$n \propto \frac{1}{\frac{A.r}{E}}$$

$$n \propto \frac{E}{A.r}$$

n.Ar = C.E

m = C.E

$$m \propto E$$

QUESTION;

In an experiment, 1930c liberated 0.64g copper. When the same quantity of electricity was passed through a solution containing silver ions ($Ag^+$), what amount of Ag was deposited. Explain the results

$M_{Ag}$ =?

$M_{cu}$ = 0.64g

$$E_{cu} = \frac{64}{2} = 32$$

$$E_{Ag} = \frac{108}{1} = 108$$

$$M_{Ag} = \frac{E_{Ag}}{E_{cu}} \times M_{cu}$$

$$M_{Ag} = \frac{108}{32} \times 0.64$$

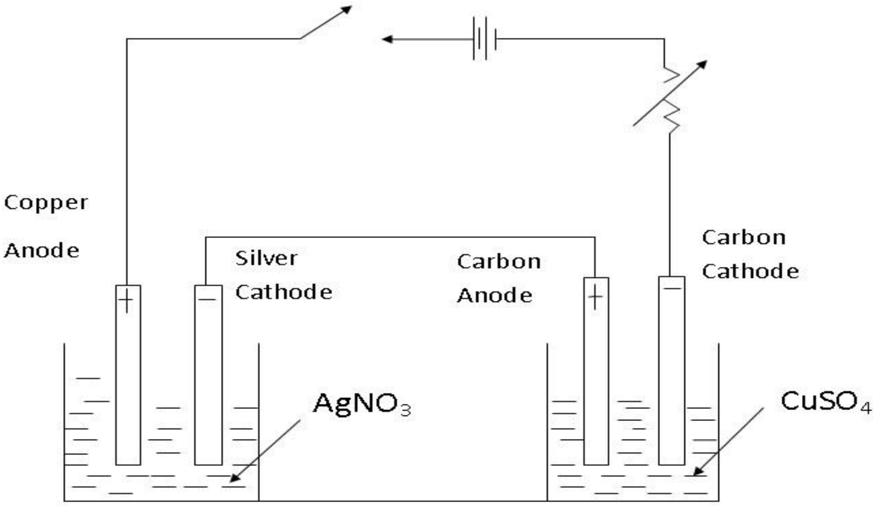
$M_{Ag}$ = 21.6g

Mass of silver is 21.6g

Mass of silver is large than of copper because silver has a small chemical equivalent than that of copper.

Mathematically Faraday’s second law;

$$\frac{m_1}{m_2} = \frac{E_1}{E_2}$$



| | |
|---|---|
| $n \propto 1/v$ | $m \propto E$ |

**QUESTIONS:**

1. A silver voltmeter is connected in series with copper voltmeter. What weight of Ag will be deposited in one voltmeter while 0.16g of copper is deposited in the other?

2. The same current passes through solution of same concentration of $PgNO_3$ and $CuSO_4$ liberated 0.23 of Ag of the same current

Chemical equivalent of Cu = 3.18 and Ag = 108.

3. A current was passed through a voltammeter arranged in series. If the weight of copper on the electrode of cells no 1 was 1.27.Find the weight of Ag deposited on the cathode of cell number 2, if same quantity of electricity is passed through

Chemical equivalent of Ag = 108

Cu = 31.75

4. An electric current was passed in series through solutions of calcium chloride and CuSO4. Carbon electrodes were used in both electrolytes. If 2.5l of chlorine measured at S.T.P were produced, what volume of oxygen would also be produced? What mass of copper was produced?

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Solution

1. $M_{cu} = 0.16$

$M_{Ag} = ?$

$E_{cu} = 64 \times \frac{1}{2} = 32$

$E_{Ag} = 108/1 = 108$

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

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

$$M_2 = \frac{0.16 \times 108}{32}$$

$M_{Ag} = 0.54g$

1kg = 1000g

? = 0.54g

$$= \frac{0.54g \times 1kg}{1000g} \times 10kg/N$$

$W_{Ag} = \underline{0.0054N.}$

2. $M_{Ag} = 0.23g$

$E_{cu} = 31.8$

$E_{Ag} = 108$

$M_{cu} = ?$

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$$\frac{M_{Ag}}{M_{Cu}} = \frac{E_{Ag}}{E_{Cu}}$$

$$\frac{0.23g}{X} = \frac{108}{31.8}$$

$$\frac{108x}{108} = \frac{31.8 \times 0.23g}{108}$$

$$x = \frac{3.657}{54} = 0.068$$

$M_{Cu}$=0.068g.

3. $M_{Cu}$ = 1.27g

$M_{Ag}$ =?

$E_{Ag}$ = 108

$E_{Cu}$=31.75

$$\frac{M_{Ag}}{M_{Cu}} = \frac{E_{Ag}}{E_{Cu}}$$

$$\frac{1.27g}{x} = \frac{31.75}{108}$$

$$\frac{31.75x}{31.75} = \frac{108 \times 1.27g}{31.75}$$

X = 4.3g

$M_{Ag}$ = <u>4.3g.</u>

4. $CaCl_2 \rightarrow Ca^{2+} + 2Cl^-$

$H_20 \rightarrow H^+ + 0H^-$

To find Faraday's

$$Cl^- \rightarrow Cl + e^-$$

$$Cl^- \rightarrow Cl + e^-$$

$$\overline{2Cl^- \rightarrow Cl_2 + 2e^-}$$

$$x = \frac{2.5\ x\ 2F}{22.4l}$$

$X = 0.223F$

X = 0.223F

Since same current is passed;

1. Volume of oxygen

$$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$$

4F = 22.4dm$^3$

0.223F =?

$$= \frac{0.223F\ x\ 22.4dm3}{4F}$$

= 1.25dm$^3$

Mass of Cu produced?

$$Cu^{2+} + 2e^+ \rightarrow Cu$$

$$2F \rightarrow 64g$$

$$0.223F \rightarrow ?$$

$$= \frac{0.223F \times ?}{2F}$$

= 7.14g

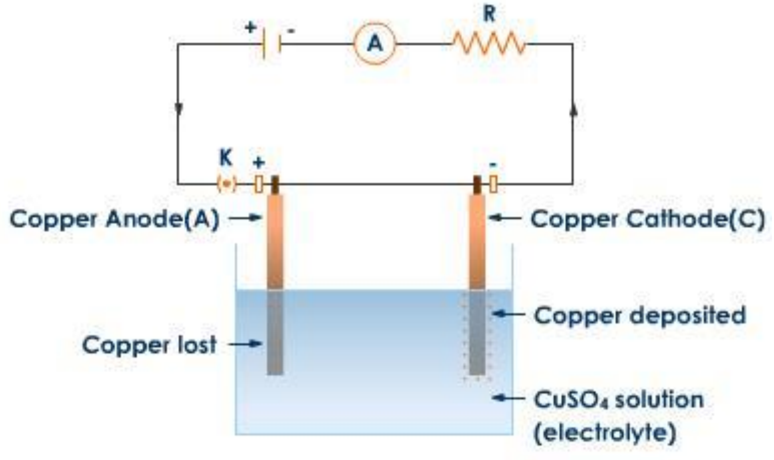
## APPLICATION OF ELECTROLYSIS

### 1. PURIFICATION OF METALS

In the method blocks of the impure metal are made at the anode and sheets of extremely pure metal are made at the cathode

Example; Copper is industrially purified by the electrolysis of aqueous copper II sulphate using copper electrodes where anode is impure and cathode is pure.

Also Zn is purified by electrolysis of Zn sulphate



### 2. MANUFACTURE (EXTRACTION) OF ELEMENTS

Very reactive elements are extracted by electrolysis of their molten chlorides e.g. Al, Mg, Na, K (Details in extraction of metals)

There is chemical combination

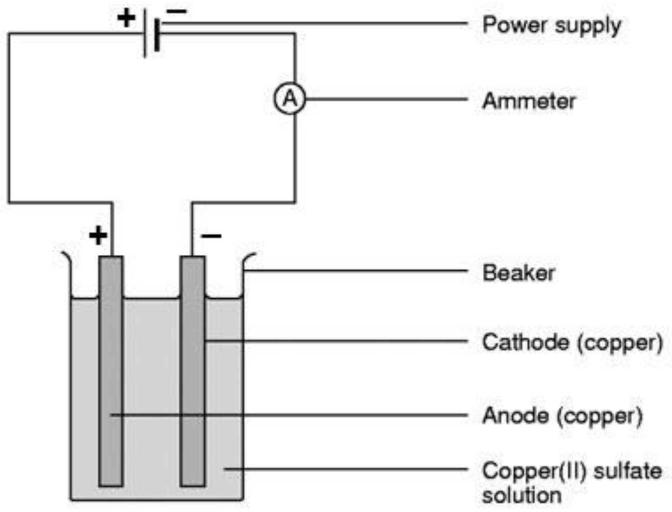
Example Au is less reactive this can be extracted without electrolysis is process

### 3. ELECTROPLATING

This is the process of covering the corroding material (metals) by the unreactive metals using electrolysis to prevent corrosion and improve their appearance.

The article to be plated is made the cathode and the anode is made of the planting material

Example planting chromium with $Ag^+$



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## CHEMICAL KINETICS, EQUILIBRIUM AND ENERGETICS

## RATE OF REACTION EQUILIBRIUM AND ENERGETICS

### 1. RATE OF CHEMICAL REACTION

This is the speed at which reaction is proceeding per unit time

$$\text{Slope of the graph} = \frac{\Delta P}{\Delta t}$$

It will be positive because product is increasing. The slope gives the rate of reaction

The SI unit = mol $dm^{-3}S^{-1}$

If the plot is **reactant concentration** against **time** the **slope will be negative**

**NB:** The units of rate depend on the units of reactant or products used

**Example;** In an experiment of determining rate, 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 shown on the syringe against time

b) Determine the rate of reaction

## 2. FACTORS AFFECTING RATE OF REACTION

a) Temperature

b) Concentration

c) Catalyst

d) Surface area

e) Pressure (gases)

f) Light

### A) **TEMPERATURE**

The increase in temperature increases the rate of chemical reaction. This is because when temperature increases the particles acquire kinetic energy, hence move faster.

For this matter the collision of the reacting particles also increases.

For every $10^0$c rise the reaction rate doubles

### B) **CONCENTRATION**

The reaction occurs when the reacting substances are allowed to come into contact. The rate of reaction therefore depends on collision frequency which in turn depends on how crowded (concentrated) the particles are

### C) SURFACE AREA

The size of particles in heterogeneous reaction influences the rate of chemical reactions. The smaller the particles, the larger the surface area and thus faster rate of reaction

Example; lumps of marble chips ($CaCO_3$) react slowly with dilute hydrochloric acid while the powder of $CaCO_3$ reacts fast with dilute HCl

Aluminium foil slowly reacts with aqueous solution of NaOH when warmed together compared to aluminium powder

### D. LIGHT

Light energies some of the reacting molecules (particles). In such reactions, molecules absorb energy in form of light energy rather than heat energy and so they are called **photochemical reactions**.

Examples are;

1. Photosynthesis

$$6CO_2 + 12H_2O \xrightarrow[\text{Chlorophyll}]{\text{light}} C_6H_{12}O_6 + 6O_2 + 6H_2O$$

2. Decomposition of silver halides on exposure to light

3. Photography

Silver bromide (AgBr) is used in photographic film. It decomposes an exposure to light as

$$AgBr \xrightarrow{\text{light}} 2Ag_{(s)} + Br_{2(s)}$$

When light falls on the film, it causes the decomposition of AgBr as shown above. The extent of decomposition depends on the brightness of sun light.

Different amounts of light reflected from the dark and light surfaces darken the AgBr film to different degrees producing a photograph.

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## E. **CATALYST**

A catalyst can either slow down or speed up the rate of reaction.

### HOW DO CATALYSTS SPEED UP THE RATE OF REACTION?

For the substance to react they must overcome the energy barrier (Activation energy, $E_A$)

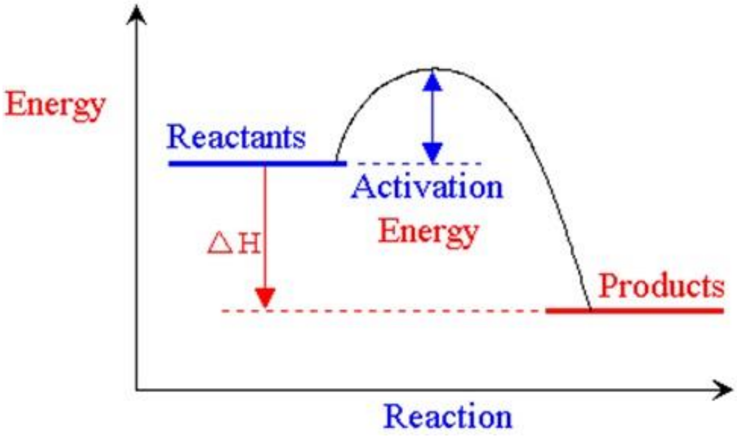
If $E_A$ is too great, the substances cannot react at room temperature because the particles don't process enough energy to surmount the barrier. In such cases, the supply energy in form of heat enables the reaction to proceed

$$\text{Rate} \propto \frac{1}{E_A}$$

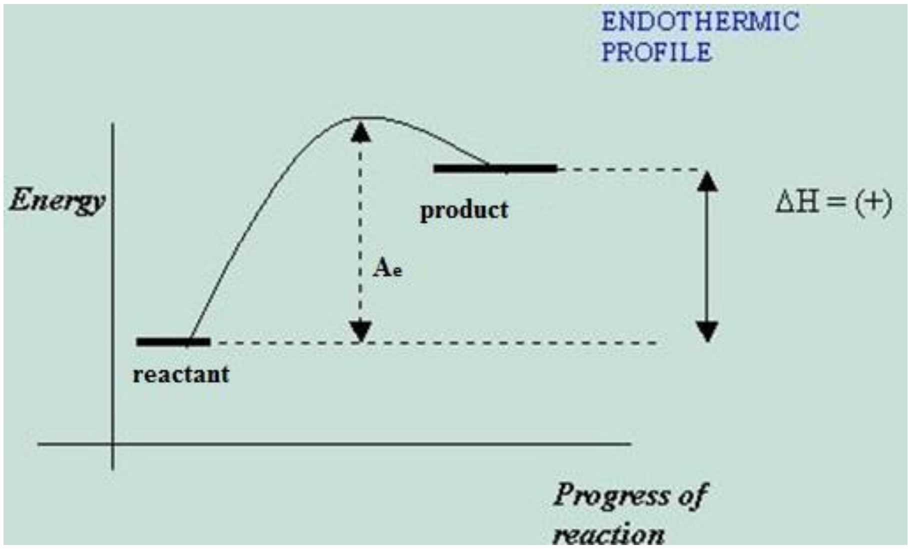
1. **Exothermic reactions** – release heat to the surrounding

2. **Endothermic reaction** – Absorb heat from the surrounding

Exothermic reaction



Endothermic reaction



**EXPLANATIONS.**

1. The energy relationship between reactants and products of a chemical reaction.

A catalyst increases reaction rate by lowering activation energy which is a sort of "Energy hump", thus providing an easier path from reactants to products.

2. **PRESSURE**

(For reactants in gaseous phase)

The increase in pressure brings the gaseous particles closer and closer together. This in turn increases the collision frequency. The latter increases the rate of reactions.

3. **<u>REVERSIBLE AND IRREVERSIBLE REACTIONS</u>**

**Reversible reactions** are reactions which proceed in both directions (background)

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

$$I_2 + H_2 \rightleftharpoons 2HI$$

**IRREVERSIBLE reactions** are the reactions which proceed to forward direction only regardless of conditions applied

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$$CaCO_3 \xrightarrow[\triangle]{} CaO + CO_2$$

$$2KClO_3 \xrightarrow[\triangle]{MnO} 2KCl + 3O_2$$

## 4. CHEMICAL EQUILIBRIUM

**Chemical equilibrium** is the state at which the forward and backward reactions are proceeding at the same rate.

That is there's some sort of balance between the reactants and products

Chemical equilibrium occurs in a **CLOSED SYSTEM.**

$$NH_4Cl \underset{\text{Cool}}{\overset{\triangle}{\rightleftharpoons}} NH_3 + HCl$$

Consider the analogous system of compounds in the same phase

$$nA + mB \rightleftharpoons xC + yD$$

As soon as little C and D are formed a reverse reaction will begin. At the beginning the forward reaction will predominate but as C and D accumulate the reverse reaction builds up until an equilibrium is reached (The forward and reserve reaction proceed at the same rate)

The point at which the rate of forward reaction = the rate of backward reaction is called **the equilibrium point or position.**

**Le Chatelier's principle**:

States that: "If a chemical system in equilibrium and one of the factors involved in the equilibrium is changed, the equilibrium will shift so as to minimize the effect of the change"

### FACTORS AFFECTING CHEMICAL EQUILIBRIUM

### A. TEMPERATURE

The effects of temperature in chemical equilibrium depends on whether chemical reaction is exothermic or endothermic

a) Exothermic reaction give out heat to the surrounding ($\Delta H$ = -ve)

b) Endothermic absorbed heat from the surrounding ($\Delta H$ =+ve)

In **exothermic** reactions

$$nA + mB \rightleftharpoons xC \quad -\Delta H$$

When heat is supplied the equilibrium will shift to the direction that requires more heat

I.e. backward reaction

In other words, the effect of heat will be on the forward reaction, thus to minimize the effect the equilibrium must shift backwards.

In **endothermic reaction**

$$pA + qB \rightleftharpoons yC \ + \Delta H$$

The increase in temperature will pose stress on the backward reaction (LHS). So to minimize the effect, the equilibrium must shift to the forward reaction

B. **PRESSURE**

Pressure affects the equilibrium of reversible reactions involving gases. The increase in pressure causes the equilibrium to shift so as to produce substances occupying less volume or (moles)

Example;

$$\underbrace{2NO_{(g)} + O_{2(g)}}_{\text{3 vol}} \rightleftharpoons \underbrace{2NO_{2(g)}}_{\text{2 vol}}$$

Increase in pressure will favour the forward reaction (position of equilibrium will shift to right) as it is accompanied by a decrease in volume. Decrease in pressure will favour backward reaction i.e. position of equilibrium will shift to the left

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For the gaseous reaction in which there is no change in volume, pressure has no effect on position of equilibrium

Example; $\underbrace{I_2 + H_2}_{2\text{ vol}} \rightleftharpoons \underbrace{2HI}_{2\text{ vol}}$

Hence if pressure is increased, the position of equilibrium will shift in such a way as to reduce the number of moles present.

The pressure exerted by gases is directly proportional to the number of moles present

C. **CONCENTRATION**

If concentration of one substance is increased, the reaction will move in a direction to use up the substance whose concentration is increased

Example; $3Fe_{(s)} + 4H_2O_{(g)} \rightleftharpoons Fe_3O_{4(s)} + 4H_{2(g)}$

What will be the effect of adding more steam to the POE – position of equilibrium of this reaction?

Adding more steam will pose an effect in the equilibrium of the backward reaction. Thus to minimize the effects the reaction must proceed forward therefore the POE will shift to the RHS.

NB: A reaction in which a gas is evolved and allowed to escape or in which a precipitate is formed will proceed in one direction only, even though a reversible reaction might be expected

Example 1; $FeCl_{3(aq)} + H_2S_{(g)} \rightleftharpoons 2FeCl_2 + 2HCl_{(aq)} + S_{(s)}$

$S_{(s)}$ precipitate, it’s almost completely removed from the system. Backward reaction is then not possible

There is complete conversation of reactants to product

Example 2; $2H_2O_{(l)} \rightleftharpoons H^+ + OH^-$

$H^+$ are removed as $H_{2\,(g)}$ Thus water dissociation has to keep going on

The forward reactions will continue without the backward reactions thus equilibrium is affected

D. **CATALYST**

A catalyst doesn't upset the equilibrium of the system because both reactions will proceed at a faster rate

A catalyst doesn't affect the POE

1. Define the following i) Exothermic reaction

   ii) Endothermic reaction

b) State the Le Chatalier's principle

c) The equation for dissociation of calcium carbonate is

$$CaCO_3 \rightleftharpoons CaO + CO_2 \ldots\ldots\ldots\ldots\ldots \Delta H = +175.5\ KJmol^{-1}$$

What will be the effect on the proportion of calcium carbonate in the equilibrium mixture if

i) Temperature is increased

ii) Pressure is increased

What are the necessary conditions for manufacture of calcium oxide from calcium carbonate in large scale?

i) **Increase in temperature** the forward reaction is favoured, that is $CaCO_3$ will be splitting to form CaO and $O_2$ thus the effect on the proportion of $CaCO_3$ in the equilibrium will be the decrease of $CaCO_3$.

ii) **Increase in pressure**

Pressure only affects substances in equilibrium in gaseous phase. Thus, $CaCO_3$ and CaO are considered to have zero volume since they do not react with pressure, but since $CO_2$ reacts with pressure its volume is considered to be 1. Therefore, backward reaction will be favoured thus the effect on the proportion will be the increase in $CaCO_3$

**Conditions**

- High temperature
- Low pressure
- Large surface area means $CaCO_3$ should be in powdered form
- Concentration

  By removing $CO_2$ from the reaction , the concentration of CaO in the products side will favour more quicker forward reaction to replace the lost $CO_{2(g)}$ thus more CaO will be manufactured.

2. Define the following;

i) Reversible reaction

ii) Rate of chemical reaction

iii) Catalyst

b) Bahati attempted to prepare hydrogen gas by reaction between Zn metal with $H_2SO_4$ .In this experiment, Zn metal of about 0.5cm diameter was used and 0.2 moles of acid. The rate of formation of hydrogen gas was found to be slow. Explain 3 ways in which the rate of formation of hydrogen could be increased.

c) If bahati wanted 36dm$^3$ of hydrogen at S.T.P. What amount of zinc would be required if 0.2moles of acid were used

The equation for the reaction is;

$$Zn_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$$

b) Answer;

i. Decreasing the diameter of zinc so as to increase the surface area.

A large diameter of zinc (o.5cm) yields less amount of hydrogen, thus in order for large amount of hydrogen to be produced, the zinc diameter has to be reduced

ii. Introduce catalyst

iii. Increase concentration of $H_2SO_4$
Thus faster reaction and rate.

iv. Increase temperature

Thus increase K.E of the particles so as to taste up the reaction.

## 5. ENDOTHERMIC AND EXOTHERMIC REACTION.

Exothermic and endothermic reactions are the types of reactions which associate internal energy change to external energy of surrounding environment.

A coldness and hotness in the environmental surrounding after stopping reaction is a result of internal change of energy for reactant to product.

### Internal Energy

Internal energy is that energy stored by molecules of chemical substances. When chemical reaction takes place of course influenced by energy changes. The internal energy decreased as a result of giving out, and also the internal energy increases as a result of inlet.

### EXOTHERMIC

An exothermic reaction is one during which heat is liberated to the surrounding.

**Example;** Combustion reaction

The heat change notation represented by change in heat. If quantity heat mentioned in (-ve) sign, Internal energy decreased and that is exothermic reaction.

**e.g** i. $H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(l)}$; $\Delta H = -286KJmol^{-1}$

ii. $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$; $\Delta H = -406KJmol^{-1}$

### ENDOTHERMIC

An endothermic reaction is one during which heat is absorbed from the surrounding.

**Example;**

A reaction to form liquid of carbon disulphide which used to manufacture Acetylene gas. Acetylene gas applied in the welding cylinders.

If the quantity heat mentioned by (+ve) sign that is increasing of internal energy. The increasing of internal energy is Endothermic reaction.

**e.g** i. $\frac{1}{2}N_{2(g)} + \frac{1}{2}O2_{(g)} \rightarrow NO_{(g)}$; $\Delta H = {}^{+}90.3KJmol^{-1}$

ii. $C_{(s)} + 2S_{(s)} \rightarrow CS_{2(l)}$; $\Delta H = {}^{+}117KJmol^{-1}$

### ENERGY LEVEL DIAGRAMS

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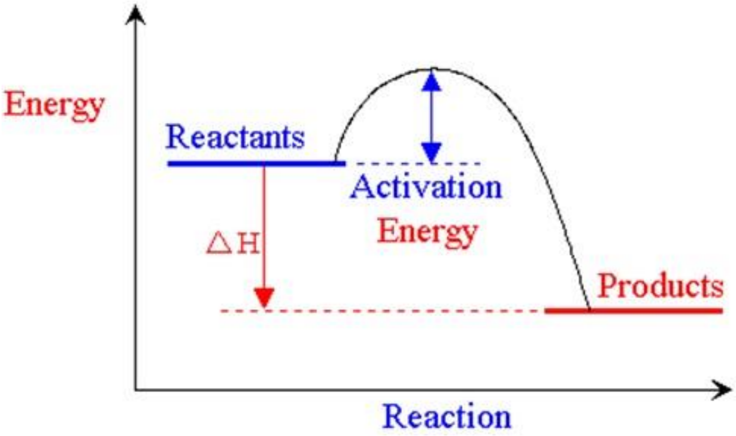
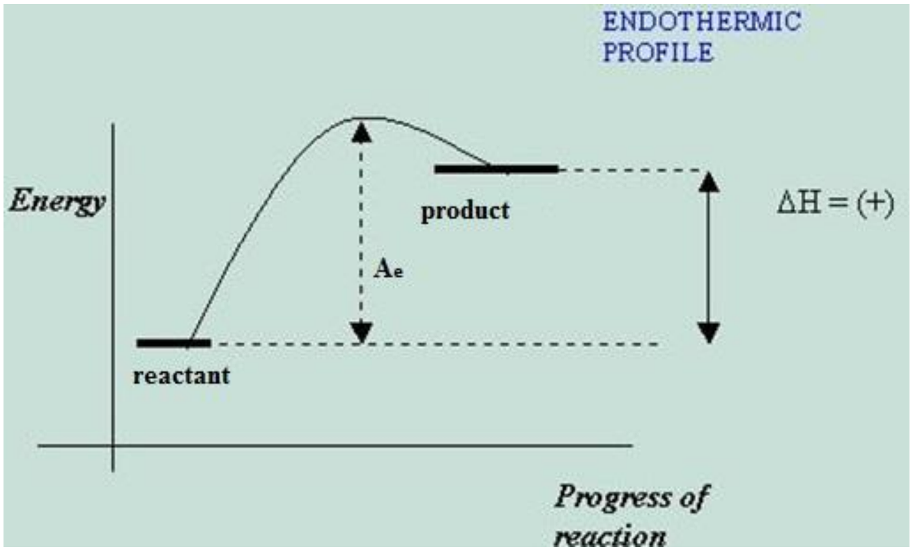
### i. EXOTHERMIC



### ii. ENDOTHERMIC



Where by Ea = Is Activation energy.

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## HARDNESS OF WATER

### INTRODUCTION.

Water is very essential substance for all living things. Over 70% of the earth is water and there are different sources of water for daily use. These sources include rivers, lakes, wells, ponds, springs, and streams.

Water can also be obtained from taps that are supplied with treated from the local or nearest water works.

In some places, water lathers easily with soap, this water is said to be **soft**. In other places, the same amount of soap would give scum and very little lather. This water is said to be **hard**. Scum is an insoluble layer of impurities that accumulates at the surface of a liquid, especially water.

Hardness of water is caused by dissolved calcium and magnesium compounds. They include calcium, calcium hydrogen carbonate, magnesium sulphate and magnesium hydrogen carbonate. For example

Calcium sulphate + sodium stearate → calcium stearate + sodium sulphate

(Hardness) (Soap) (Scum)

The soap and the hardness join to form scum. The other compound sodium sulphate which is formed dissolves in water.

Calcium hydrogen carbonate is the most common cause of hardness of water. It forms when rain falls on the rocks of limestone and chalk, which are mainly composed of the insoluble calcium carbonate. As the rain falls through the air, it dissolves carbon dioxide to form weak acidic solution. This solution is able to attack the calcium carbonate to form the soluble calcium hydrogen carbonate.

Water + carbon dioxide + calcium carbonate → calcium hydrogen carbonate

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

Dolomite and gypsum are other rocks that are sources of hard water.

## TYPES OF HARDNESS OF WATER.

There are two types of hardness of water, namely Temporary and Permanent hardness of water.

Hardness caused by calcium hydrogen carbonate is called temporary hardness because it can be removed by boiling the water.

$$Ca(HCO_3)_{2\ (aq)} \rightarrow H_2o_{\ (l)} + CO_{2\ (g)} + CaCO_{3\ (s)}.$$

Hardness caused by other calcium and magnesium compounds is called **Permanent Hardness.**

This is because boiling does not affect it.

## HOW TO REMOVE PERMANENT HARDNESS OF WATER

Permanent hardness of water can be removed by:

**i) Distillation;** This gets rid of both temporary and permanent hardness.

**ii) Adding sodium carbonate (washing soda);** This is added to the water to precipitate calcium carbonate. It removes both types of hardness of water for example, its reaction with sodium sulphate is;

Calcium sulphate+ sodium carbonate →calcium carbonate

$$CaSO_{4\,(aq)} + Na_2CO_{3(aq)} \rightarrow CaCO_3(s) + Na_2SO_{4\,(aq)}$$

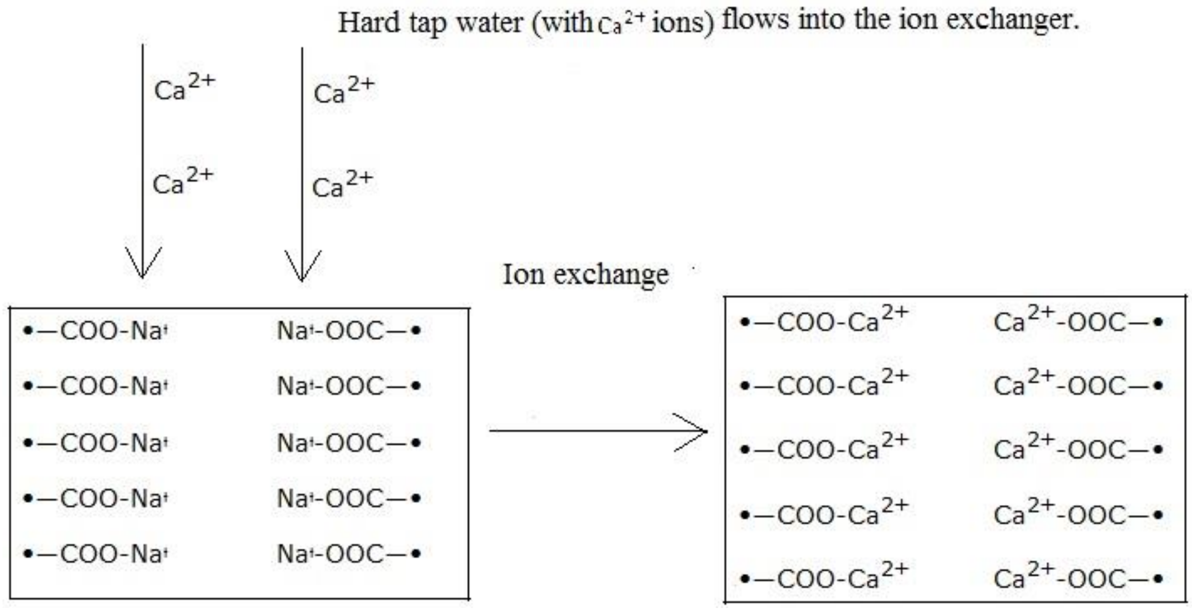
**iii) Ion exchangers;** These can remove both types of water hardness by removing all the calcium and magnesium ions in the water.

An ion exchanger is a container full of small beads. These are made of special plastic called **ion exchange resin.** This has ions for example, sodium ions that are weakly attached to it.

When hard water is passed through the ion exchanger, the calcium and magnesium ions in the water change places with the sodium ions and attach themselves to the resin. The calcium and magnesium ions are therefore left behind in the resin as the soft water flows out with the sodium ions.

After some time, all the sodium ions will have been replaced and the exchanger cannot remove hardness anymore. This is resolved by regeneration of the resin by pouring a concentrated solution of sodium chloride into the exchanger.

The sodium ion push the calcium and magnesium ions off the resin, making the ion exchanger ready for use again.



The $Ca^{2+}$ from the hard water replaces the $Na^{+}$ ions in the ions exchanger; softened water (with $Na^{+}$ ions) leaves the ion exchanger to be used in the household.

The sample of water that do not lather easily are of hard water. These include the sea water and sometimes the tap water. The boiled, distilled and rain water easily form lather with soap and are therefore soft water.

## ADVANTAGES AND DISADVANTAGES OF HARD WATER.

### 1) ADVANTAGES:

i) It tastes better due to dissolved compounds.

ii) It provides useful calcium for the growth of bones and teeth.

iii) The formation of lime scale in pipes forms a sort of insulation which prevents the water in the tap from coming into contact with the metal of the pipe. This prevents pipe corrosion and poisonous metal salts from getting dissolved in the water.

### 2) DISADVANTAGES:

i) The temporary hardness of water causes limescale in water boilers, hot water pipes, kettles and other appliances, this makes them less efficient.

Scaling can also cause blockage in the appliances and has to be removed from time to time.

ii) Hard water needs more soap than soft water, so it means a lot of soap is wasted.

iii) Hard water leaves scum deposits that are difficult to wash out and causes damage various articles.

## WATER TREATMENT.

Before water is used in the public water supply, it has to be treated, not to make it pure but to make it safe to drink.

The various stages include:

1. **Filtration** to remove any large objects such as algae or parasites.

2. **Precipitation** of mineral salts such as iron salts using alum (calcium hydroxide).

3. **Filtration** through a sand bed- this is a biological process. A gelatinous layer is used ton trap unwanted particles.

4. **Chlorination.** Treating with chlorine kills harmful bacteria in the water- this is a disinfectant.

5. **Fluoridation**. Adding sodium fluoride in small quantities. Fluorides harden the enamel on teeth and reduce dental problems. Not all water is treated in this way.

As much as a supply of clean, safe water is essential to ensure good health, we must not forget that safe disposal of used water is of equal importance. This becomes especially important when people live close together in towns and cities.

In a city it is essential to have a system for the disposal of used water. We call this a sewage system.

The processes involved in sewage treatment plants are outlined below.

1. The sewage that enters a treatment plant contains debris that might damage the pumps and machinery such materials are removed by screens or vertical bars. After removal, this debris is burned or buried.

2. The waste water then passes through a grinder where leaves and other organic materials are reduced in size for efficient treatment and removal later.

3. Next, grit is removed in grit chambers. The grit is removed and the disposed of as sanitary landfill.

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4. With grit removed, the water passes into a sedimentation tank, in which organic materials settle out and are drawn off for disposal. An alternative to sedimentation that is used in the treatment of some waste water is flotation, in which air is forced into the waste water under pressure of 1.75 to 3.5kg mâ□»$^2$. The waste water, supersaturated with air, is then discharged into an open tank, there the rising air bubbles cause the suspended solid to rise to the surface, where they are removed.

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## EXTRACTION OF METALS

### OCCURRENCE OF METALS.

The metals which are most in reactivity are extracted from the sea, for example; calcium is extracted from limestone, chalk and marble in the sea.

- Metals of medium reactivity are found in the form of oxides and sulphides

- Example; Aluminium, Zinc, Iron and Tin

- Least reactive metals are found as free uncombined elements, for example; Gold, Silver etc.

### Main natural/forms of metals.

| Metals in order of reactivity | Main form in nature |
|---|---|
| Potassium | potassiumchloride |
| Sodium | sodium chloride |
| Calcium | calcium carbonate which is limestone / chalk |
| Aluminium | aluminium oxide |
| Zinc | zinc sulphide |
| Iron | iron oxide or iron sulphide |
| Tin | tin oxide |
| Lead | lead sulphide |

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| | |
|---|---|
| Copper | copper pyrites ($CuFeS_2$) |
| Mercury | mercury sulphide |
| Silver | silver sulphide |
| Gold | occurs as a free element |
| Platinum | occurs as a fee element |

## LOCATION OF METALS IN TANZANIA

Tanzania is a country of mineral rich land.A varieties of metals investigated to exist in many parts of Tanzania. A mass medias of Tanzania are very familiar to some important places like Nyamongo, Buzwagi, Bulyanhulu etc. Those places are very important for the economical fluctuation and personal financial enhancement. Many people migrate there, as a means solving their way of life.

| METALS | LOCATION |
|---|---|
| 1. Copper | - Morogoro |
| 2. Tin | - Misemyi Kagera |
| 3. Phosphate | - Minyingu in the lake Narton Manyara |
| 4. Uranium | - Namtumbo Ruvuma |
| 5. Iron Haematite ore | - Liganga place<br>- Ludewa district Njombe |
| 6. Gold | - Chunya Mbeya<br>- Nyarugusu Geita<br>- Nyamongo Tarime district, Mara<br>- Bariadi<br>- Buzwagi and Bulyahulu Shinyanga |

## EXTRACTION OF METALS

Extraction is the process whereby metals are obtained from their ores. The ores are reduced to respective metals in the process. The metallic ions gain electrons and become the corresponding atom

Example; $Na^+ + e^- \rightarrow Na$

$$Cu^{2+} + 2e^- \rightarrow Cu$$

$$Fe^{3+} + 3e^- \rightarrow Fe$$

-The extraction can be done in two methods;

1. Electrolytic reduction
2. Chemical reduction

Table shows some metals and their methods of extraction

| | **METAL** | **NAME OF ORE** | **EXTRACTION METHOD** |
|---|---|---|---|
| Very reactive metals | Potassium | Pota – chloride | Electrolysis |
| | Sodium | Rock salt (NaCl) | |
| | Magnesium | Magnesium chloride | |
| | Aluminium | Bauxite | |
| Less reactive metals | Zinc | Zinc Blender | Chemical reduction |
| | Tin, Iron | Hematite, magnetite | |
| | Lead | Galena | |
| | Copper | Copper pyrites | |
| Least reactive metals | Gold | Uncombined | Purified by electrolysis |

## STAGES OF EXTRACTION OF METALS.

There are two stages in the process of extraction of metals from their ores;

1. Ore purification

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2. Extraction of the metal
3. Refining of the metal

## 1. ORE PURIFICATION

After mining the ore from particular process, the ore should be purified. There are almost three methods that can be used to purify the ore.

a) Dressing
b) Calcination
c) Roasting

a) Dressing

Is the method used to remove impurities without decomposing any component within the ore. Impurities like sand , limestone, quartz and silicate should be removed by;

- Hand picking
- Blowing
- Filtrating solids
- Sedimentation and Decaution

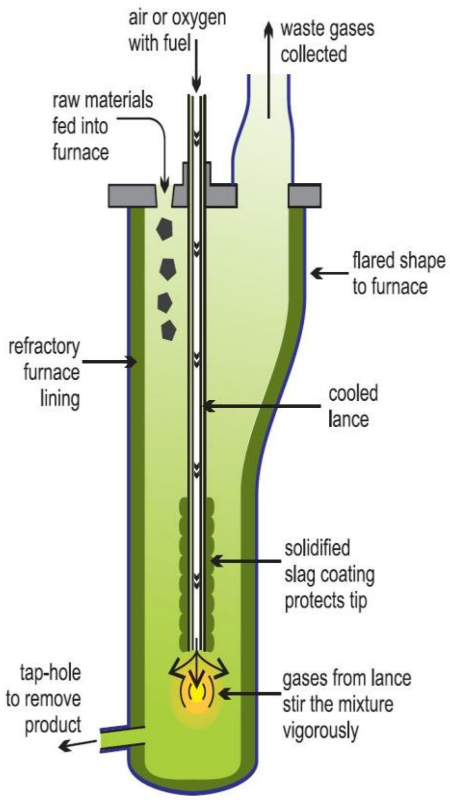
- Dressing also known as concentration.

b) Calcination

Calcination is a process in which ore is heated in the absence of air (to avoid oxide product) below its melting point to expel water from a hydrated oxide or expel carbondioxide from a carbonate.

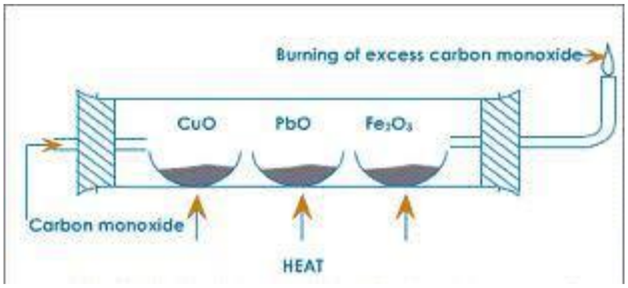
c) Roasting

Roasting is a method used to purify ore by heating in the presence of air. Some time the ore may be mixed with other chemicasl. The roasting needed to make great change to oxide or chlorination.



## 2. REDUCTION OF OXIDES TO METALS

Carbon reduces oxides of zinc (Zn) and all metals below Zinc (Zn) in the E.C.S metals above Zinc (Zn) in the E.C.S require very high temperature for reduction. So many metals are obtained through electrolysis.



## EXTRACTION OF SODIUM

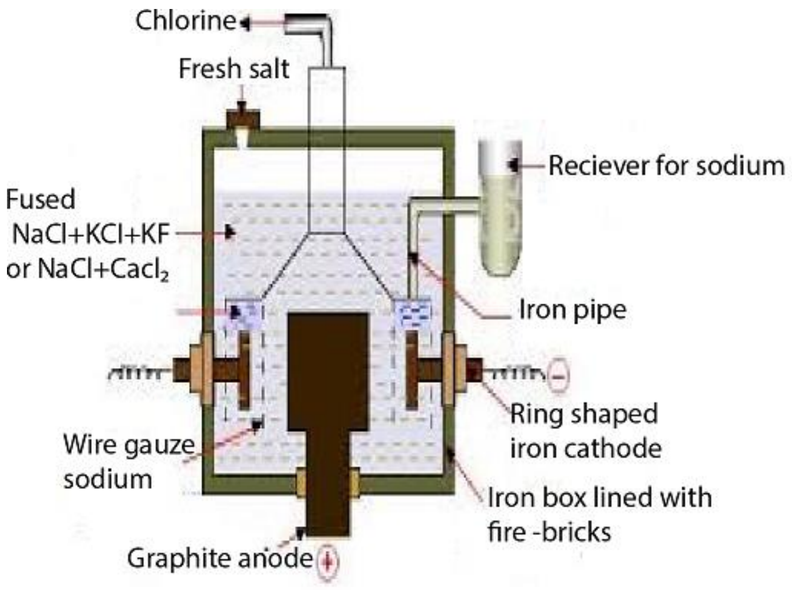
Sodium is extracted by electrolysis of a mixture of molten sodium chloride and calcium chloride. The calcium chloride is added in order to reduce the temperature from over $800^0$c to $600^0$c.

The cell used is Down's cell, sodium is liberated as the cathode, while chlorine is liberate at the anode

At the cathode $Na^+ + e^- \rightarrow Na$

At the anode $2Cl^- \rightarrow Cl_2 + 2e^-$

The two elements are separated from each other by cylindrical iron gauze



## EXTRACTION OF ALUMINIUM

Aluminium is extracted by electrolysis from its ore (oxide $Al_2O_3$) which occur as hydrated aluminium oxide (bauxite) $Al_2O_3.2H_2O$.

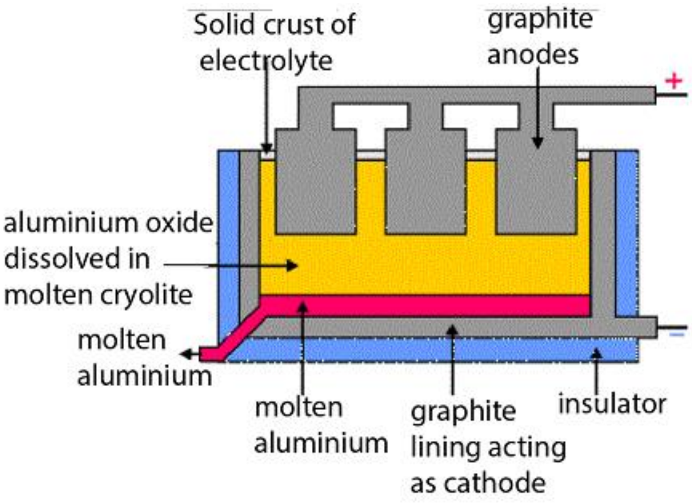
The purified bauxite is dissolved in molten cryolite ($Na_3AlF_6$) and is electrolyzed. The anode and cathode are both carbon

At cathode

$Al^{3+} + 3e^- \rightarrow Al$

At anode

$2O^{2-} \rightarrow O_2 + 4e^-$



## USES OF ALUMINIUM

1. it's used for making kitchen vessels such as pots, pans etc.

2. it's used in moving parts of machines and in engines example; pistons and cylinders.

3. it's used for making objects that must be as light as possible

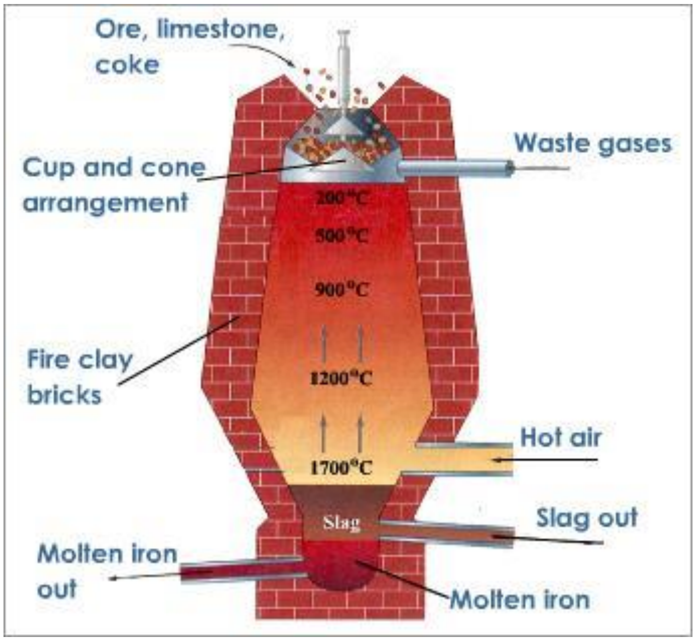
Example parts of aero planes, railways trains and truck, buses, lorries, tankers, furniture and car etc

4. in over head tension cables for distribution of electric power

5. Other uses are packing materials for cigarette, sweets, biscuits etc.

## EXTRACTION OF IRON AND STEEL

The main ores of iron are haematite, $Fe_2O_3$ and Magnetite $Fe_3O_4$, iron II carbonate, $FeCO_3$ also occurs. The ore is heated to expel water. It is then loaded into the top of blast of furnace together with coke and limestone



Iron is extracted from its ore by reduction in a blast furnace

STAGE 1;

Near the base of the furnace where a hot air blast enters by the tuyeres, coke burns in air to produce carbon dioxide, providing the main source of heat for the process

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$$

STAGE 2;

A little higher in the furnace the carbon dioxide is reduced by hot coke to carbon monoxide

STAGE 3;

At higher stage still carbon monoxide reduces the ore to iron, some reduction may also occur directly by carbon

$$Fe_2O_{3(s)} + 3CO_{(g)} \rightarrow 2Fe_{(s)} + 3CO_{2(g)}$$

$$Fe_2O_{3(s)} + 3C \rightarrow 2Fe_{(s)} + 3CO_{(g)}$$

As it settles in the furnace the iron melts and drops into the bottom of the funnel

## THE FUNCTION OF THE LIMESTONE ($CaCO_3$)

The limestone is heated and decomposes quick lime (CaO) and $CO_2$

$$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$$

The quicklime (CaO) reacts with acid impurities in the ore to form molten slag

$$CaO_{(s)} \rightarrow SiO_{2(s)} + CaSiO_{3(s)}$$

Slag and iron are tapped off separately at frequent intervals. The iron being cast into "pigs" and is known as **pig iron** or **cast iron**. It contains many impurities like C, P, Si, Mn, S and thus it is brittle. It can however be cast by pouring into moulds as liquid to produce many articles in which great strength is necessary, Si, Mn, S and thus it is brittle. It can however be cast by pouring into moulds as liquid to produce many articles in which great strength is necessary.
**Example;** Railings, water pipes, bases of Bunsen, cookers, stoves etc

## WROUGHT IRON:

It is the purest form of commercial iron. It contains 99% iron and less than 0.25% of carbon, to make wrought iron, cast iron is melted and stirred in a furnace carbon in the cast iron is oxidized away of gas (CO and $CO_2$) while other impurities
**Example**; Mn and Si are oxidized to a slag

$$Fe_2O_{3(s)} + 3C_{(s)} \rightarrow 2Fe_{(s)} + 3CO_{(g)}$$

**Function of limestone- to remove impurities**

(i) The wrought iron is strong and malleable. It can be shaped by hammering at very high temperature, $1000^0$c.

(ii) It is used for making nails, sheets, chains, gates, farm machines etc.

## STEEL

This is an alloy of iron with 0.15% to 0.17% or carbon. Steel is obtained by oxidizing the impurities in molten pig iron using a jet or "tonnage oxygen" directly vertically down the metal; after removing of impurities a calculated mass of carbon and (Ni, Co, Cs, Mn, It needed) is added.

## PROPERTIES OF METALS

### PHYSICAL PROPERTIES OF METALS AS COMPARED TO NON – METALS

| PROPERTY | METALS | NON – METALS |
|---|---|---|
| Thermal or electrical conductivity | High | Low |
| Lustre | High | Low |
| Sonority | High | Low |
| Ductility | High | Low |
| Tensile strength | High | Low |
| Malleability | High | Low |

### 1. Metals are lustrous

Can be polished and also can be rolled into different shapes. That is why they are useful in manufacture of light weight cans.
Example; Aluminium can be rolled into very thin foils which can be used even for rapping sweets. Lead is soft and bendable. It's useful in protecting underground cables. Because of its ductility and malleability steel can be brought into different shapes

**NB**: Malleable substance can be hammered into different shapes. It can be flattened into sheets

- Ductile substance can be drawn into a wire
- A lustrous substance is shiny and can be polished by a steel wool to make it shine.
- A sonorous substance can make noise when hit

### 2. Metals conduct heat

Metals are used to make cooking utensils like kettles and saucepans because they are good conductors. Heat can be conducted from source to the food being cooked through them.

### 3. Metals have high melting points

Although not all metals have high melting points (example Na and K have low densities and float on water and their melting points are low; Na 98°C K 63°C and also mercury which is a liquid metal mp -39°C) most metals do.

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The use of tungsten in filaments; alloys of chromium and nickel in heating elements of electric fires and engine in cars and aero planes do utilize this property.

**4. Metals have high tensile strength**

Metals can support heavy loads steel is useful in making girders, hawsers and chains because of this property the head of hammer is useful because it's hard and tough.

## CHEMICAL PROPERTIES OF METALS

**1. Physical strength and chemical strength**
- Physical strength results from the way atoms are arranged in a substance, example; Metals of high tensile strength like Fe, Cu and Al and low tensile strength like Na and K.
- Chemical strength, we look in reactivity of the element. This depends on how the electrons are arranged in atom. K and Na have got very low tensile strength but chemically they are the strongest metals.

**2. Reducing power of metals**
- The reactivity depends on the easiness to donate electrons i.e. reducing power. A more reactive metal can displace a less reactive metal from its compound.
- For example; when we roast copper (i) sulphide we get Cu while when we roast lead sulphide, no lead is made

I) $Cu_2S + O_{2(g)} \rightarrow 2Cu_{(s)} + SO_{2(g)}$

ii) $2PbS_{(s)} + 3O_{2(g)} \rightarrow 2PbS_{(s)} + 2SO_{2(g)}$

**NB:** The more reactive a metal is, the more difficult it is to extract it from its ore

**3. Displacement reactions of metals and reactivity**
A more reactive metal can displace a less reactive metal from an aqueous solution or its oxide.

For example; i) $Mg + ZnO \rightarrow MgO + Zn$

Magnesium + zinc oxide

ii) $Zn_{(s)} + PbNO_{3(aq)} \rightarrow ZnNO_{3(aq)} + Pb_{(s)}$

Zinc + Lead nitrate

### 4. Reaction of metals with water

**-** As reducing agents, metals react with water to liberate $H_{2(g)}$. K and Na react violently with water while Ca and Mg react violently with steam and Al, Zn, Fe, Pb and Cu have no action on water

### 5. Reaction of metals with dilute HCl.

- The reaction between K, Na and Ca and dilute HCl to liberate $H_{2(g)}$ is violent and dangerous, while Mg, Al Zn, Fe and Sn and Pb give off $H_{2(g)}$ very slowly.

-The acid needs to be warmed up Cu, Ag and Au are not attached at all.

---

## COMPOUNDS OF METALS

### A. METAL OXIDE

An oxide is a compound of oxygen with another element.

**Preparation of oxides:**

There are two ways of preparing metal oxides

$$\left.\begin{array}{l}\bullet\ \text{Direct}\\ \bullet\ \text{Indirect}\end{array}\right\}\ \text{Methods}$$

### a) DIRECT METHOD

It involves direct combination of metal with oxygen

Example; $2Ca_{(s)} + O_{2(g)} \rightarrow 2CaO_{(s)}$

$$2Mg_{(g)} + O_{2(g)} \rightarrow 2MgO_{(s)}$$

$$4Na_{(s)} + O_{2(g)} \rightarrow 2Na_2O_{(s)}$$

### 1. INDIRECT METHOD

It involves thermal decomposition of a salt of carbonate, hydroxide (I) nitrates

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Example; $CaCO_{3(s)} \underset{\Delta}{\rightarrow} CaO_{(s)} + CO_{2(g)}$

$2Pb(NO_3) \underset{\Delta}{\rightarrow} 2PbO + NO_2 + O_2$

$2Cu(NO_3) \underset{\Delta}{\rightarrow} 2CuO + NO_2 + O_2$

## CLASSIFICATION OF METAL OXIDES

### 1. BASIC OXIDES

Are oxides of metal when reacts with acid forms salts

Examples are MgO, CaO, $Na_2O$, $K_2O$

$$MgO + 2HCl \rightarrow MgCl_2 + H_2O$$

$$CaO + H_2SO_4 \rightarrow CaSO_4 + H_2O$$

### 2. ACIDIC OXIDES

Are oxides of non metals and when dissolved in water form solution

Examples are $CO_2$, $NO_2$, $SO_2$

$$CO_2 + H_2O \rightarrow H_2CO_{3(\text{carbonic acid})}$$

$$NO_2 + H_2O \rightarrow H_2NO_3$$

### 3. AMPHOTERIC OXIDES

These are oxides which have both basic and acidic properties

(H neutralize both acidic and Alkali to form salt)

Examples are ZnO, $Al_2O_3$ and PbO

### 4. NEUTRAL OXIDES

Are oxides which do not react with either acid or base (Have no basic/acidic properties)

Examples are $H_2O$, CO, NO, $N_2Cl$

5. **PEROXIDES**

When elements burn in excess air, peroxides are formed

Examples are $N_2O_2$, $H_2O_2$

6. **MIXED OXIDES**

These are oxides of two simple oxides

Examples are $Fe_3O_4$ – Tri iron tetra oxide

(Mixture of Fe O and $Fe_2O_3$)

**SOLUBILITY.**

| | |
|---|---|
| K<br>Na | Oxides of this metal are soluble in water |
| Ca | Oxides of calcium are sparingly soluble in water |
| Mg<br>Al<br>Zn<br>Fe<br>Cu | Oxides of this metal are insoluble in water |

## USES OF OXIDES

I. Used in preparation of salts in the laboratory

I.e. $CaO + H_2SO_4 \rightarrow CaSO_4 + H_2O$

II. Used in formation of slag

III. Used as a drying agent

IV. Used in manufacture of motors

V. Oxides react with acid to form salt and water

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Example; $ZnO + 2HCl \rightarrow H_2O + ZnCl_2$

The following are the reaction equations when the following oxides react with hydroxides

i) $ZnO + 2NaOH \rightarrow H_2O + Na_2ZnO_2$(Sodium zincate)

ii) $Al_2O_3 + 2NaOH \rightarrow H_2O + 2NaAlO_2$(Sodium aluminate)

iii) $PbO + 2NaOH \rightarrow H_2O + Na_2PbO_2$(Sodium plumbate)

B. **METAL HYDROXIDE**

Very active metals (K, Na and Ca) dissolves in water forming hydroxide soluble oxides dissolves in water to form hydroxide (Alkali)

Soluble oxides dissolve in water to form hydroxide (alkali)

Examples; $Na_2O + H_2O \rightarrow NaOH$(Caustic soda)

$K_2O + H_2O \rightarrow KOH$(Caustic potash)

$CaO + H_2O \rightarrow Ca(OH)_2$(Lime water slaked lime)

Caustic soda and caustic potash are used in the manufacture of **SOAP.**

Lime water is used to detect $CO_2$ gas. The rest of the hydroxides are insoluble.

## PREPARATION OF INSOLUBLE HYDROXIDES

i) Dissolve a soluble salt in water i.e. $CU(N0_3)_2$

ii) Dissolve a soluble hydroxide in water i.e. $Ca(OH)_2$

iii) Mix them together

$Cu(NO_3)_2 + Ca(OH)_2 \rightarrow Cu(OH)_2 + Ca(NO_3)_2$

iv) Separate the two by filtration.

NB: The filtrate is $Ca(NO_3)_2$ and the residue is $Cu(OH)_2$

## PROPERTIES OF HYDROXIDE

i) Have bitter taste

ii) Neutralize acids to form salt and water

$$HNO_3 + NaOH \rightarrow NaNO_3 + H_2O$$

Solubility of hydroxide in water

- The nature of hydroxide of metals varies according to position of the metal in the reactivity series.

Example;

K
Na
Ca
Hydroxide of these metals is soluble in water

Mg
Al
Zn
Fe
Pb
Cu
Hydroxides of these metals are insoluble in

Hg
Au
Ag
Hydroxide of these metal do not exist

The hydroxide of sodium, calcium and potassium are the strongest because they dissociate completely in water

Example; $KOH \rightarrow K^{+} + OH$

Ammonium hydroxide is weak base because it does not dissociate completely in water

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$$NH_4OH \rightarrow NH_4^+ + OH$$

**Question:**

By using diagram and details explain/ show how you would prepare

i) Zinc hydroxide

ii) Magnesium hydroxide

**Solution:**

i) Zinc hydroxide ($Zn(OH)_2$)

- Dissolve soluble salt in water ($Zn(NO_3)_2$).

- Dissolve soluble hydroxide in water (NaOH).

- Mix them together.

$$Zn(NO_3)_2 + 2NaOH \xrightarrow{\qquad\qquad} Zn(OH)_{2(s)} + NaNO_{3(aq)}$$

Separate by filtration

## CARBONATE AND HYDROGEN CARBONATE

- Carbonates occur in many natural forms examples are chalk, limestone and marbles. Examples of carbonates that occur naturally are the carbonates of Zinc, iron, lead and Manganese.
- Sea animals have carbonates in their shells in the forms of calcium carbonates. The remains of these animals sink to the bottom of sea leading to chalk formation after thousands of years, through pressure the chalk hardens to form limestone

**General properties of carbonates:**

1. All carbonates are insoluble in water except those of sodium, potassium and ammonium.

2. All carbonates give carbon dioxide on heating except those of sodium and potassium

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For example; $CuCO_3 \rightarrow CuO + CO_2$

(Green) (Black)

3. All carbonates give carbon dioxide with dilute acids

For example; $ZnCO_3 + H_2SO_4 \rightarrow ZnSO_4 + CO_2 + H_2O$

## PREPARATION OF SODIUM/ POTASSIUM CARBONATES AND SODIUM/ POTASSIUM HYDROGEN CARBONATES:

a) A solution of sodium hydroxide (caustic soda) is saturated with carbon dioxide, giving a solution of sodium bicarbonates

$$NaOH + CO_2 \rightarrow NaHCO_3$$

If a solution of potassium hydroxide (caustic potash) is saturated with carbon dioxide, a solution of potassium bicarbonate is formed

$$KOH + CO_2 \rightarrow KHCO_3$$

a) This solution is divided into two equal portions, one portion is allowed to evaporate at room temperature, when crystals of sodium hydrogen carbonate (sodium bicarbonate) or potassium hydrogen carbonate (potassium bicarbonate) are obtained which are be filtered and dried.

b) An equal volume of the original sodium hydroxide or potassium hydroxide solution is added to the second portion. Sodium carbonate or potassium carbonate is formed

$$NaHCO_3 + NaOH \rightarrow Na_2CO_3 + H_2O$$

$$KHCO_3 + KOH \rightarrow K_2CO_3 + H_2O$$

The solution is evaporated down and allowed to cool crystals of sodium carbonate is formed

## DIFFERENCES BETWEEN SODIUM CARBONATE AND SODIUM BICARBONATE

**($Na_2CO_3$)**

**($NaHCO_3$)**

| | | |
|---|---|---|
| 1 | Have 2 crystalline forms $Na_2CO_3.10H_2O$ and $Na_2CO_3$ and anhydrous form sodium carbonate efflorescence. | - Can only be made in the hydrous from sodium hydrogen carbonate does not efflorescence ($NaHCO_3$) |
| 2 | Is soluble in water | Is much soluble in water |
| 3 | Does not decompose on heating | It decomposes on heating to form carbonate, carbon dioxide and water<br>$2NaHCO_3 \rightarrow Na_2CO_3 + CO_3 + H_2O$ |
| 4 | Is known as washing soda used in softening water and removing grease | Sodium hydrogen carbonate is used as a baking powder to make bread, rise $CO_2$ given off causes the bread of cake to rise |

All carbonates and hydrogen carbonates react with acids to form carbon dioxide, water and salt.

For example; $K_2CO_3 + H_2SO_3 \rightarrow K_2SO_4 + CO_2 + H_2O$

Dilute sulphuric acid react very little with calcium carbonate because the product calcium sulphate is insoluble, form a layer on the outside of the carbonates and stop further reaction.

Also dilute sulphuric acid and hydrochloric acid react little with lead II Carbonate because the product, the lead sulphate and lead II Chloride are insoluble, they form layer on the outside of the carbonates which cause to stop further reaction

- Nitric acid react readily with all carbonates because all nitric are soluble.

## TEST FOR CARBONATES AND HYDROGEN CARBONATE.

a) Add dilute nitric acid to a solution of a substance. Carbon dioxide gas is evolved, water and salts are formed.

Examples; i) $CaCO_3 + 2HNO_3 \rightarrow CO_2 + H_2O + Ca(NO_3)_2$

ii) $NaHCO_3 + HNO_3 \rightarrow CO_2 + H_2O + NaNO_3$

b) i) Add magnesium sulphate to a substance, if carbonates are present white precipitates are formed.

Examples; $Na_2CO_3 + MgSO_4 \rightarrow MgCO_3 + Na_2SO_4$

ii) If hydrogen carbonates are present no precipitates are formed

c) **Action of heat**

Potassium carbonate and sodium carbonate do not decompose when heated. Other carbonate decompose to form an oxide and carbon dioxide

a) $PbCO_3 \underset{\Delta}{\rightarrow} CO_2 + PbO$

b) $ZnCO_3 \underset{\Delta}{\rightarrow} CO_2 + ZnO$

c)$MgCO_3 \underset{\Delta}{\rightarrow} CO_2 + MgO$

d) $CuCO_3 \underset{\Delta}{\rightarrow} CuO + CO_2$

e) $CaCO_3 \underset{\Delta}{\rightarrow} CO_2 + CaO$

**Preparation of sodium carbonate by SOLVARY PROCESS**

1. Calcium carbonate is heated in a limekiln

$$CaCO_{3(s)} \underset{\Delta}{\rightarrow} CaO_{(s)} + CO_{2(g)}$$

2. The calcium oxide is dissolved in water

$$CaO + H_2O \underset{\Delta}{\rightarrow} Ca(OH)_{2(aq)}$$

3. The carbon dioxide is reacted with ammonia brine

$$2NaCl_{(aq)} + 2NH_{3(g)} + 2H_2O_{(l)} + 2CO_2 \rightarrow 2NH_4Cl_{(aq)} + 2NaHCO_{3(aq)}$$

As the solution cools, sodium hydrogen carbonates the least soluble combination of ions, precipitates out and removed

4. The sodium hydrogen carbonate is heated

$$2NaHCO_3 \underset{\Delta}{\rightarrow} Na_2CO_3 + H_2O$$

5. The Ammonium chloride and calcium hydroxide are reacted

$$Ca(OH)_{3(aq)} + NH_4Cl_{(aq)} \rightarrow CaCl_{(aq)} + 2NH_{3(g)} + 2H_2O_{(l)}$$

## NITRATES

Nitrates are important chemicals for industries purpose

Examples;

- potassium nitrate is used in gun powder
- Sodium nitrates occur in nature as salt petres
- Aluminium nitrate is useful in the manufacture of sulphuric acid as a fertilizer.

## PREPARATION OF NITRATES

1. By neutralizing nitric acid with an alkali or base

2. by the action of nitric acid on a carbonate

3. by the action of metal on nitric acid copper

(Only for magnesium, zinc, lead and copper)

a) $NaOH + HNO_3 \rightarrow NaNO_3 + H_2O$

b) $CaCO_3 + HNO_3 \rightarrow Ca(NO_3)_2 + H_2O + CO_2$

c) $Mg + 2HNO_3 \rightarrow Mg(NO_3)_2 + H_2$

## PREPARATION OF NITRATES

1. All nitrates are soluble and form crystalline solids

2. Action of heat on nitrates

a) Sodium and potassium nitrates give nitrites

$$2NaNO_3 \rightarrow 2NaNO_2 + O_2$$

(Sodium nitrite)

$$2KNO_3 \rightarrow 2KNO_2 + O_2$$

(Potassium nitrite)

b) Many metal nitrates give the oxide

Example; i) Effect of heat on copper nitrate crystals when the blue crystals of copper nitrates are heated in a boiling tube they dissolve in their water of crystallization, prolonged heating gives brown fumes of nitrogen dioxide and the residue, copper oxide is a black.

$$2Cu(NO_3)_2 \rightarrow 2CuO + 2NO_2 + O_2$$

(Blue) (Black) (Brown colorless)

ii) Effect of heat on lead nitric crystals

- When heated they decrepitate (expand with a crackling noise) and melt. Prolonged heating gives brown fumes of nitrogen dioxide and a yellow residue of lead monoxide (litharge)

$$2Pb(NO_3)_2 \rightarrow 2PbO + 4NO_2 + O_2$$

(Litharge yellow) (Brown colourless)

c) Silver and mercury nitrates give the metal

$$2AgNO_3 \rightarrow 2Ag + 2NO_2 + O_2$$

$$Hg(NO_3)_2 \rightarrow Hg + 2NO_2 + O_2$$

d) Ammonium nitrates give nitrous oxide

$$NH_4NO_3 \rightarrow N_2O(\text{nitrous oxide}) + 2H_2O$$

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3. The action of concentrated sulphuric acid

   - Concentrated sulphuric acid, heated with a nitrate liberates nitric acid

4. All nitrates except those of sodium and potassium are hydrated, sodium nitrate and calcium nitrates are deliquescent, and other nitrates are not.

Summary of chemical properties of Nitrates

$$\left.\begin{array}{l}\text{K}\\ \text{Na}\end{array}\right\}\text{Its nitrate decompose to form nitrite and oxygen}$$

**TEST FOR NITRATES**

1. Add concentrated sulphuric acid to the solid nitrate and warm Acid fumes (nitric acid) are involved. Add 3 – 4 copper turning and arm the mixture, the colour of the gas turns to deep brown.

2. The brown ring test is normally used to confirm the nitrates.

To a solution of nitrates are added a few drops of freshly prepared ferrous sulphate solution.

A few drops of concentrated sulphuric acid are added carefully and brown ring obtained at the junction between the liquid.

It is nitrate ferrous sulphate ($FeSO_4NO$) formed that

$$6FeSO_4 + 2HNO_3 + 3H_2SO_4 \rightarrow 3Fe_2(SO_4)_3 + 4H_2O + 2NO$$

The nitrogen monoxide then reacts with more iron (II) sulphate to give the brown compound, which appears as a brown ring

**USES OF NITRATES**

1. Potassium nitrate is used as a food preservative; it is also used to make slow burning fuses.

2. Sodium nitrate is mainly used as a fertilizer

3. Ammonium nitrate is mixed with chalk, the mixture is known as **NITROCHALK**, widely used as fertilizer.

It is also used in manufacturing explosives.

4. Silver nitrate is used to make silver bromide and silver iodides chemicals used to make photographic film.

## CHLORIDES

Chlorine can react with all metals to form chloride. Most common metals are attacked by dil. Hydrochloric acid to form chlorides

Example is Mg, Ca, Al, Zn, and Fe.

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

Those elements (metals) which are not attached by dilute carbonates or oxides of these metals react with dil.HCl hydrochloric acid to produce chloride

## PREPARATION OF CHLORIDE

- Metallic chlorides can be prepared by direct or indirect or indirect methods.

### 1. BY DIRECT METHOD

- Chloride can be prepared by direct action of chlorine on metals

I.e. $Fe + Cl_2 \rightarrow FeCl_2$

$$2Fe + 3Cl_2 \rightarrow FeCl_3$$

Iron (iii) chloride can be prepared by passing $Cl_2$ over a heated metal.

An iron wire is placed in hand glass tube and a steam of dry $Cl_2$ passes over it.

The reaction continues without further application of heat source and iron (iii) chlorides collect in a bottle as shown.

## PREPARATION OF SOLUBLE CHLORIDE:

Soluble chlorides can be prepared by mixing dilute hydrochloric acid (Dil.HCL)

i) Oxides

ii) Hydroxides

iii) Carbonates

iv) Metals

Examples: a) $CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

b) $KOH_{(s)} + HCl_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(l)}$

c) $Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

d) $MgO_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_2O_{(l)}$

Methods for preparation of soluble chloride are summarized below

K prepared by action of the oxide, hydroxide or carbonate

Na dilute hydrochloric acid

Ca

Mg Prepared by Action of the metal, oxide or carbonate on

Al Dilute hydrochloric acid

Zn

Fe

NB: Lead and silver chloride are only common insoluble chlorides

**<u>CHEMICAL PROPERTIES OF CHLORIDES</u>**

- To identify a chlorine in a liquid mixture add nitrates or silver nitrate (A white precipitate formed)

$$Pb(NO_3)_2 + HCl \rightarrow \underset{\text{(White precipitate)}}{PbCl_2} + 2HNO_3$$

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- To identify a chloride from solid mixture, add concentrated sulphuric acid

  The gas that produces thick white fumes with ammonium solution is produced

$$NaCl_{(s)} + H_2SO_{4(aq)} \rightarrow NaHSO_{4(aq)} + HCl_{(aq)}$$

$$NH_3 + HCl \rightarrow NH_4Cl(\text{thick white fumes})$$

- Mix chloride with oxidizing agent ($MnO_2$)

  A green wish yellow gas is involved

$$NaCl_{(s)} + 2H_2SO_{4(aq)(conc)} + MnO \rightarrow MnSO_{4(aq)} + Na_2SO_{4(aq)} + H_2O_{(l)} + Cl_{2(g)}$$

## METAL SULPHATES

i) **Preparation of soluble metal sulphates**

Soluble sulphates can be prepared in the laboratory by dissolving a metal, a carbonate, a hydroxide or an oxide in dilute $H_2SO_4$

Examples;

a) $Cu_{(s)} + H_2SO_{4(aq)} \rightarrow CuSO_{4(aq)} + H_2O_{(l)}$

b) $Zn(OH)_{2(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + 2H_2O_{(l)}$

c) $ZnCO_{3(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_2O_{(l)} + CO_{2(g)}$

d) $Zn + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$

2. **Preparation of insoluble sulphates**

Insoluble sulphates can be prepared by adding dilute sulphuric acid to lead or barium ions

Examples;

a) $BaCl_{2(s)} + H_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2HCl_{(aq)}$

b) $Pb(NO_3)_{2(aq)} + H_2SO_{4(aq)} \rightarrow PbSO_{4(s)} + 2HNO_{3(aq)}$

## CHEMICAL PROPERTIES

1. All sulphates react with barium chloride to give a white precipitate which is insoluble in dilute HCl

Note. Sulphites and carbonates give precipitate which is soluble in dilute HCl.

2. The sulphates or iron (II), copper (II) and zinc (called green, blue and white vitriol respectively) decompose when heated strongly to give the oxide of the metal, sulphur dioxide, sulphur trioxide and water

$$2FeSO_4.7H_2O \rightarrow FeO_3 + SO_2 + SO_3 + 14H_2O$$

Note. In case of copper (II) and Zinc Sulphates, no sulphur dioxide is produced

$$CuSO_4.5H_2O \rightarrow CuO + SO_3 + 5H_2O$$

3. The sulphates of sodium, potassium, magnesium and calcium don't decompose on heating

## USES OF SULPHATES

1. Sodium sulphate

Is used to treat wood pulp in the manufacture of paper

2. Anhydrous ($CaSO_4$) is used as a source of sulphur dioxide for the manufacture of sulphuric acid.

3. Ammonium sulphate is an important fertilizer

4. Aluminium sulphate

Is used in the purification of sewage as it precipitates proteins

5. $CuSO_4$ (copper (II) sulphate) is used as fungicide i.e. destroyer of fungi on plants

Also it is used as an electrolyte for electroplating copper.

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