CHAPTER: 1

ATOMIC STRUCTURE

ATOM

John Dalton proposed (1808) that atom is the smallest invisible particle of matter . Atomic radii are of order 10^{-8} cm. It contains three sub atomic particles namely electrons, protons and neutrons.

ELECTRON

- It was discovered by JJ Thomson.
- It carries a unit negative charge ($-1.6 * 10^{-19}$ C).
- Mass of electron is ($9.11 * 10^{-31}$ kg).

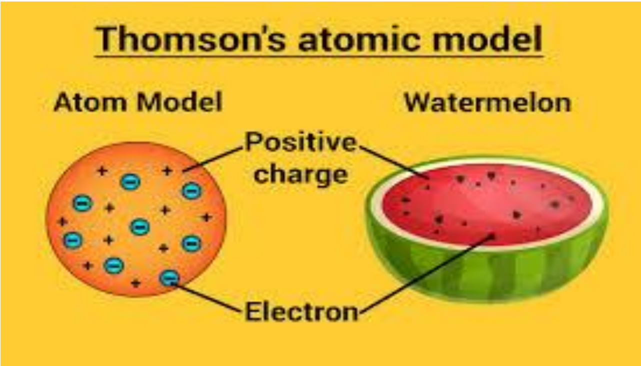
PROTON

- It was discovered by Rutherford
- It carries a unit positive charge ($+1.6 * 10^{-19}$)
- Mass of proton is 1.007276 U

NEUTRON

- It was discovered by James Chadwick
- Neutrons are neutral particles
- Mass of neutron is $1.67 * 10^{-24}$ g or 1.008665 amu

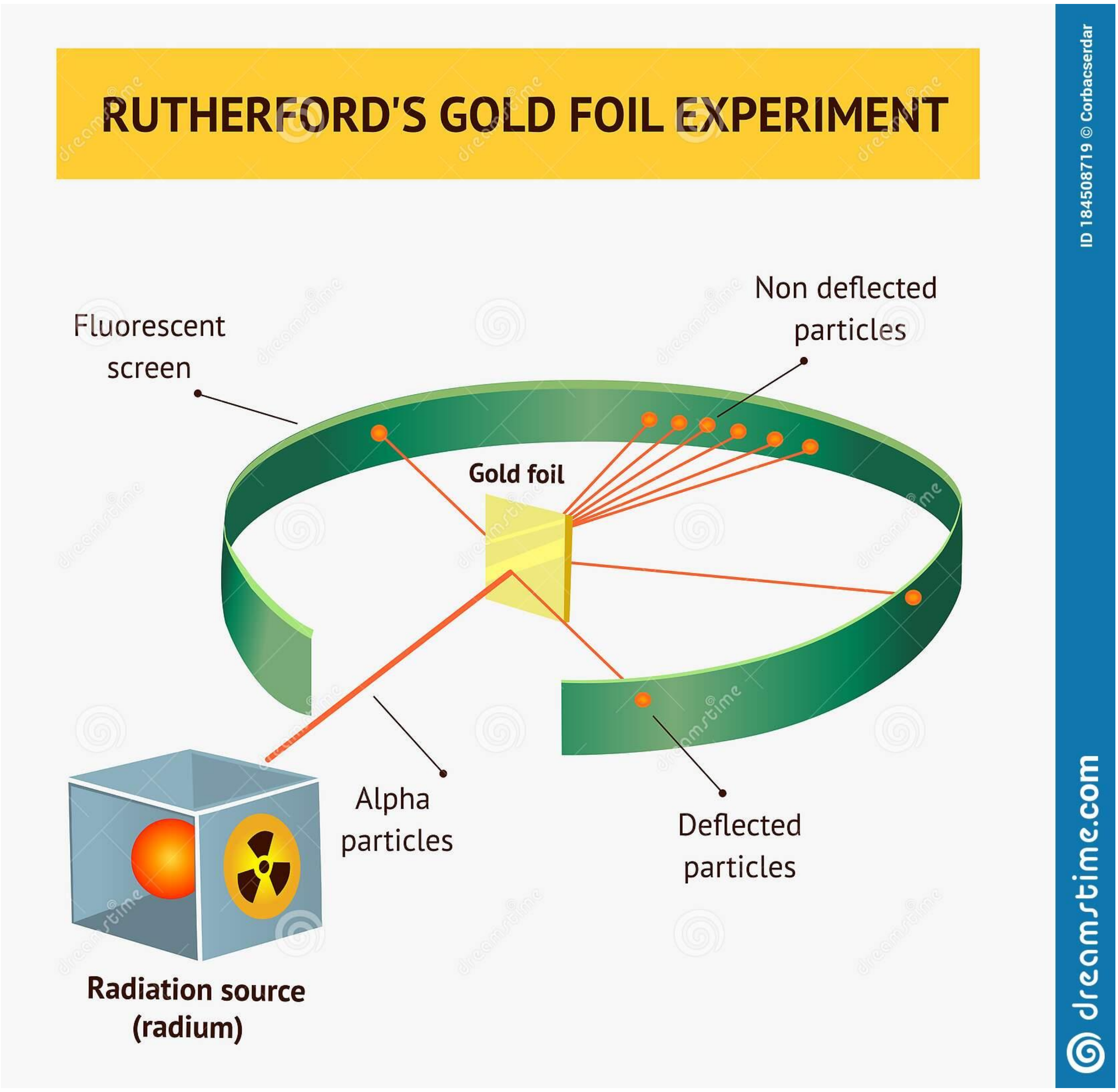
THOMSON'S ATOMIC MODEL (1898)



- According to JJ Thomson atom id a positive sphere with a number of electrons distributed within the sphere.
- It is also known as plum pudding model.
- It explains the neutrality of an atom.

- This model couldn't explain the results of ionisation and scattering experiments carried out by Rutherford.

RUTHERFORD'S ALPHA SCATTERING EXPERIMENT



- Rutherford bombarded a thin sheet of gold foil (of thickness 0.00004cm) with alpha particles.
- Alpha particles are generated from a radioactive element like radium.
- A thin lead plate with a hole cut in it was taken as a slit to form a beam of alpha particles.
- A circular screen coated with zinc sulphide was placed on the other side of the foil.

OBSERVATIONS:

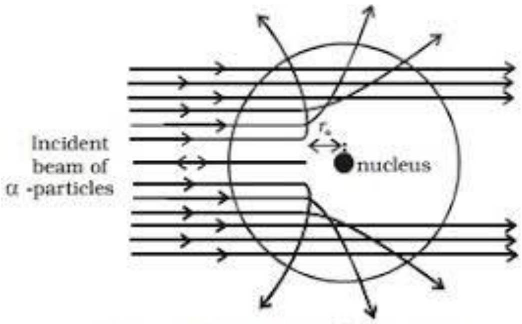
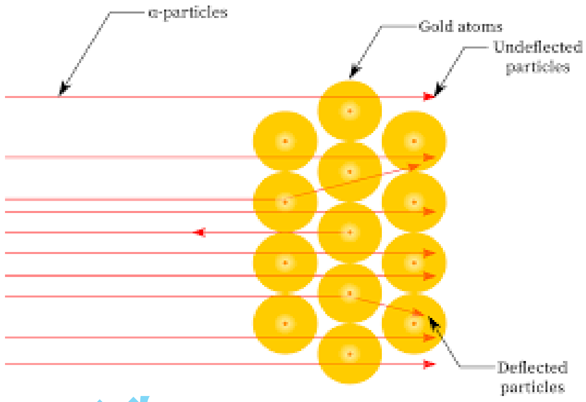


Fig Scattering of α – rays



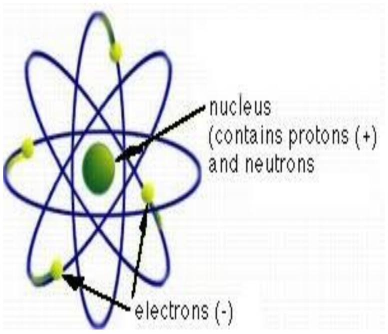
- Most of the alpha particles passed straight through the gold foil and cause illumination on the screen.
- Very few alpha particles are deflected at some angles after passing through the gold foil.
- A rare number of alpha particles (one in 10,000) were collide with gold foil and bounced back.

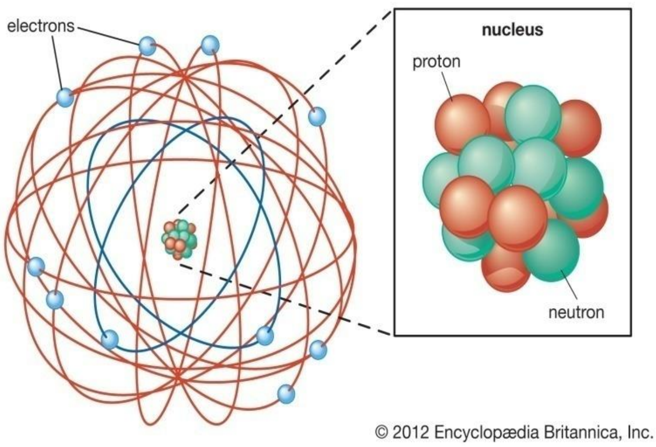
CONCLUSION:

From the above observations Rutherford concluded that

- Most of the space of an atom is empty.
- Deflection of alpha particles at some angels states that there is a small but heavy and positive charge particle is present at the centre of the atom which is called as nucleus.

POSTULATES OF RUTHERFORD'S ATOMIC MODEL

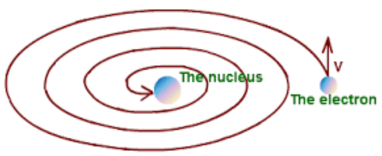




- An atom consists of two parts
  (i) Nucleus
  (ii) Extra nuclear part
- Nucleus is extremely small but heavy and positive charge situated at the centre of an atom.
- Nucleus is consist of proton and neutron.
- Nucleus carries nearly the whole mass of an atom .
- The extra nuclear part of the atom is empty space around the nucleus where electrons are revolving round the nucleus.
- As the atom is neutral the number of electrons present in it is equal to the number of protons

FAILURE OF RUTHERFORD'S ATOMIC MODEL

- It doesn't explain the distribution of electrons in an atom and their energy.
- This atomic model doesn't obey the law of electrodynamics. When a charge particle revolve round an oppositely charged centre it loses energy continuously and come closer and closer towards the nucleus and ultimately fall into the nucleus. Thus Rutherford's model of atom is faulty.



In the planetary model of atom, the electron should emit energy and spirally fall on the nucleus.

- It couldn't explain the line spectra of an atom.

POSTULATES OF BOHR'S ATOMIC MODEL

- An atom consists of a massive positively charged nucleus.
- Electron revolves round the nucleus in certain fixed circular orbits without losing or gaining energy.
- Such orbits are called stationary states or main energy levels and numbered as 1,2,3,4 etc or alphabetically designated as K,L,M,N etc respectively
- Energy associated with an orbit is given by

$$E_n = -1312/n^2 \text{ kJ/mole for H-atom.}$$

Where n = number of energy levels

- Electrons revolve only on those circular orbit for which the angular momentum is integral multiple of $h/2\pi$

$$L = mvr = nh/2\pi$$

Where, m = mass of electron , v = velocity of electron

r = radius of orbit , n = number of orbit

h = planks constant

- Transition of electrons between two stationary states can take place by absorption or emission of energy

$$\Delta E = E_2 - E_1 \text{ or } h\nu = E_2 - E_1$$

Where ΔE = change of energy
E2 = energy of second shell
E1 = energy of first shell
h = Plank's constant

- While revolving in a particular orbit an electron neither gain energy nor lose energy .The energy of a shell in an atom is fixed.

BOHR - BURY SCHEME

The distribution of electron in different orbits is given by Bohr and Bury as follows:

- The maximum number of electrons that can be present in an orbit is equal to $2n^2$.
  Where n = number of the orbit
  For 1st shell, n=1 , number of electrons = $2n^2 = 2 * (1)^2 = 2$
  For 2nd shell, n=2, number of electrons = $2n^2 = 2 * (2)^2 = 8$
  For 3rd shell, n=3, number of electrons = $2n^2 = 2 * (3)^2 = 18$
  For 4th shell, n=4, number of electrons = $2n^2 = 2 * (4)^2 = 32$
- The outermost orbit of an element cannot contain more than 8 electrons and the penultimate shell cannot contain more than 18 electrons.
- It is not always necessary to complete the orbit before the next orbit starts filling.

ATOMIC NUMBER

- Atomic number of an element is defined as the number of the protons present in the nucleus of an atom.
- Ex: Hydrogen has one proton, so atomic number of hydrogen is 1

ATOMIC MASS

- The atomic mass of mass of an element is defined as the relative average mass of its atom as compared to the mass of an atom of carbon taken as 12 (^{12}**C**).
- Hence atomic mass of an element is the number which shows how many times the mass of that atom is heavier than1/12th the mass of the carbon atom (^{12}**C**) or 1 amu.
- Atomic mass is a number and it has no unit.
- Atomic mass of nitrogen is 14 or 14 amu. That means one atom of nitrogen is 14 times heavier than 1/12th mass of one atom of carbon (^{12}**C**).

MASS NUMBER

- The sum of number of protons and number of neutrons present in the nucleus of an atom is known as the mass number.
- Ex: C has 6 protons and 6 neutrons

So mass number =6+6 =12

ISOTOPES

- Isotopes are defined as the atoms of the same element which have same atomic number but different mass number.
- Ex: $_{1}\mathbf{H}^{1}$ $_{1}\mathbf{H}^{2}$ $_{1}\mathbf{H}^{3}$
  $_{6}\mathbf{C}^{12}$ $_{6}\mathbf{C}^{13}$
  $_{17}\mathbf{Cl}^{35}$ $_{17}\mathbf{Cl}^{37}$

ISOBARS

- Isobars are defined as the atoms of different elements having same mass number but different atomic number.
- Ex: $_{18}\mathbf{Ar}^{40}$ $_{20}\mathbf{Ca}^{40}$

ISOTONES

- Isotones are defined as the atoms of different elements having same number of neutrons.
- Ex: $_{32}\mathbf{Ge}^{76}$ $_{33}\mathbf{As}^{77}$

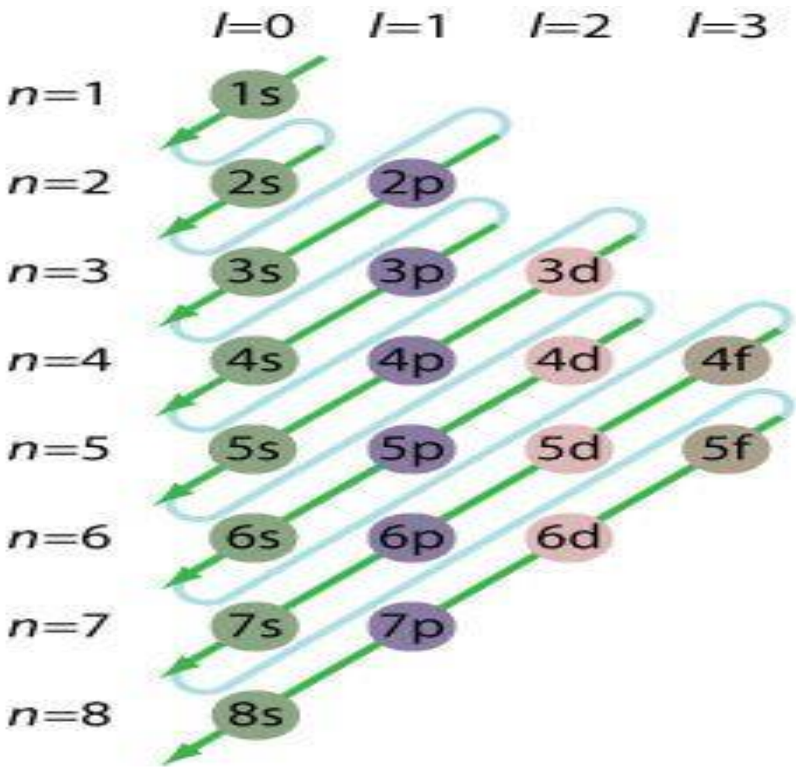
AUFBAU PRINCIPLE

- According to this principle, the electrons are filled in various orbitals in order of their increasing energies.
- Thus an orbital with lowest energy will be filled first.
- The energy of various sub-shell can be determined by (n+l) rule
- (n+l) Rule : The sub-shell with lower (n+l) value will possess lower energy and will be filled first.

Ex:

Sub-shell	(n+l) value
1s	1+0=1
2p	2+1=3
3d	3+2=5

Thus energy order of above sub-shells will be 1s<2p<3d

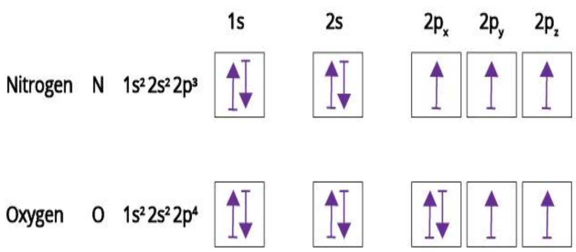


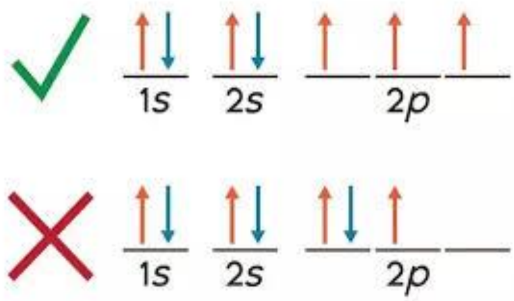
HUND'S RULE

- According to this rule, pairing of electrons doesn't take place in p, d and f orbitals unless each degenerate orbitals in the given sub-shell is single filled.

  OR

  Orbitals of same sub-shell first get single filled and then pairing occurs.
- Ex :





ELECTRONIC CONFIGURATION

Electron Configurations

Element	Orbital Filling						Electron Configuration
	1s	2s	$2p_x$	$2p_y$	$2p_z$	3s	
H	↑						$1s^1$
He	↑↓						$1s^2$
Li	↑↓	↑					$1s^22s^1$
C	↑↓	↑↓	↑	↑			$1s^22s^22p^2$
N	↑↓	↑↓	↑	↑	↑		$1s^22s^22p^3$
O	↑↓	↑↓	↑↓	↑	↑		$1s^22s^22p^4$
F	↑↓	↑↓	↑↓	↑↓	↑		$1s^22s^22p^5$
Ne	↑↓	↑↓	↑↓	↑↓	↑↓		$1s^22s^22p^6$
Na	↑↓	↑↓	↑↓	↑↓	↑↓	↑	$1s^22s^22p^63s^1$

#	Element	Electron configuration
1	Hydrogen	$1s^1$
2	Helium	$1s^2$
3	Lithium	$1s^2 2s^1$
4	Beryllium	$1s^2 2s^2$
5	Boron	$1s^2 2s^2 2p^1$
6	Carbon	$1s^2 2s^2 2p^2$
7	Nitrogen	$1s^2 2s^2 2p^3$
8	Oxygen	$1s^2 2s^2 2p^4$
9	Fluorine	$1s^2 2s^2 2p^5$
10	Neon	$1s^2 2s^2 2p^6$
11	Sodium	$1s^2 2s^2 2p^6 3s^1$
12	Magnesium	$1s^2 2s^2 2p^6 3s^2$
13	Aluminum	$1s^2 2s^2 2p^6 3s^2 3p^1$
14	Silicon	$1s^2 2s^2 2p^6 3s^2 3p^2$
15	Phosphorous	$1s^2 2s^2 2p^6 3s^2 3p^3$
16	Sulfur	$1s^2 2s^2 2p^6 3s^2 3p^4$
17	Chlorine	$1s^2 2s^2 2p^6 3s^2 3p^5$
18	Argon	$1s^2 2s^2 2p^6 3s^2 3p^6$
19	Potassium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
20	Calcium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
21	Scandium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
22	Titanium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
23	Vanadium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$
24	Chromium*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
25	Manganese	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$
26	Iron	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
27	Cobalt	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$
28	Nickel	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$
29	Copper*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
30	Zinc	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

Chapter-2

Chemical Bonding

CHEMICAL BOND

A chemical bond is defined as the force of attraction which holds the atoms together and forms molecule.

Ex: H-H   O=O   Cl-Cl   Na-Cl

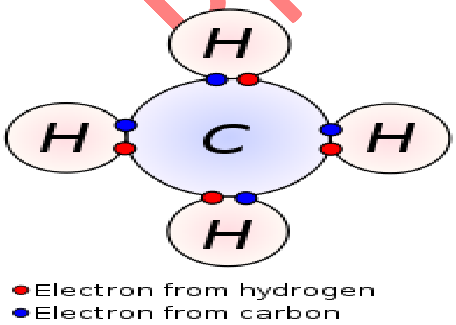
TYPES OF CHEMICAL BOND:

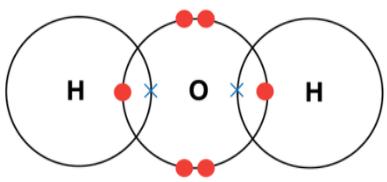
1. Covalent Bond
2. Ionic bond
3. Coordinate bond

COVALENT BOND:

It is defined as the bond which is formed by the mutual sharing of electrons between two atoms.

Ex: H_2 O_2 H_2O CH_4





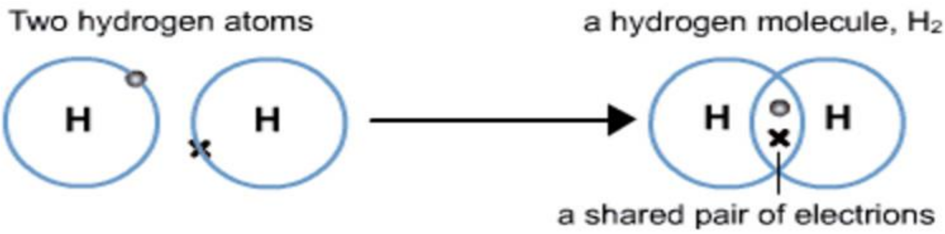
FORMATION OF H_2:

- Electronic configuration of Hydrogen is

$$_{1}\mathbf{H\text{- }1S^{1}}$$

$$_{1}\mathbf{H\text{-}1S^{1}}$$

- Two atoms of H combine to completes its duplet as Helium and becomes stable and forms H_2 molecule.



H-H or H_2

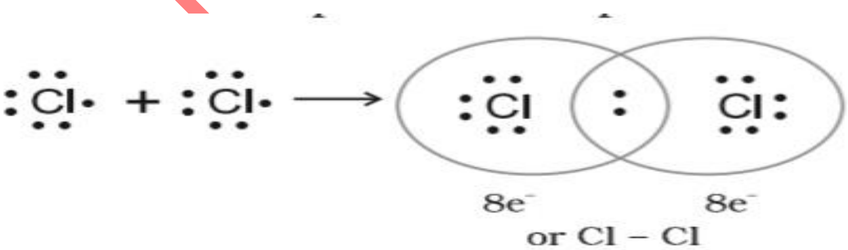
FORMATION OF Cl_2:

- Electronic configuration of **Cl** is

$$_{17}Cl: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$$

$$_{17}Cl: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$$

- Each of chlorine atom has 7 valence electrons
- Thus each chlorine atom mutually share its one electron with each other to complete its octet and forms **Cl_2** molecule.



Covalent bond between two Cl atoms

Cl-Cl or Cl_2

FORMATION OF H_2O:

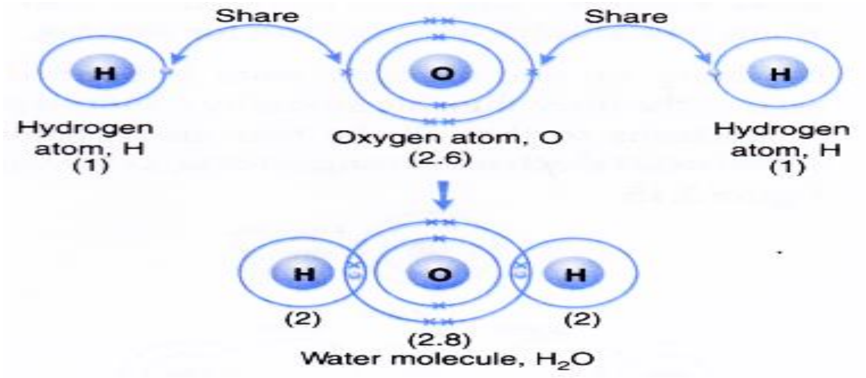
- Electronic configuration of **H** is

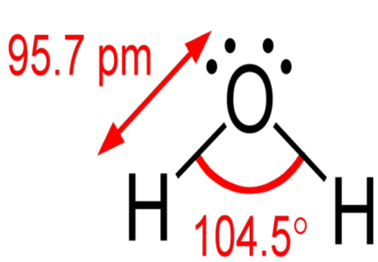
$$_{1}H\text{-}1s^1$$

Electronic configuration of **O** is

$_8O$-1s^2 2s^2 2p^4

- Hydrogen has 1 valence electron and Oxygen has 6 valence electrons.
- Thus each Hydrogen atom mutually share its one electron with an electron of oxygen to complete its octet and forms **H_2O** molecule.





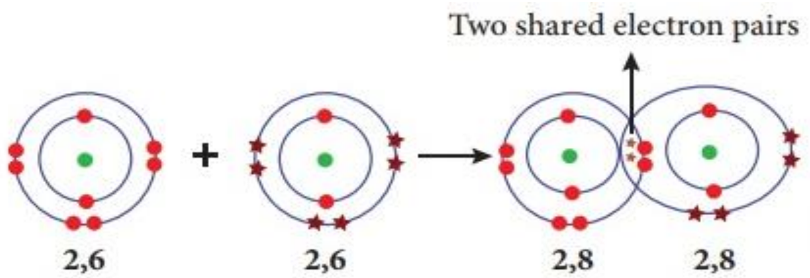
FORMATION OF O_2:

- Electronic configuration of oxygen is

$_8O$-1s^2 2s^2 2p^4

$_8O$-1s^2 2s^2 2p^4

- Each of oxygen atom has 6 valence electrons.
- Thus each oxygen atom mutually share its two electrons with each other to complete its octet and forms **O_2** molecule.



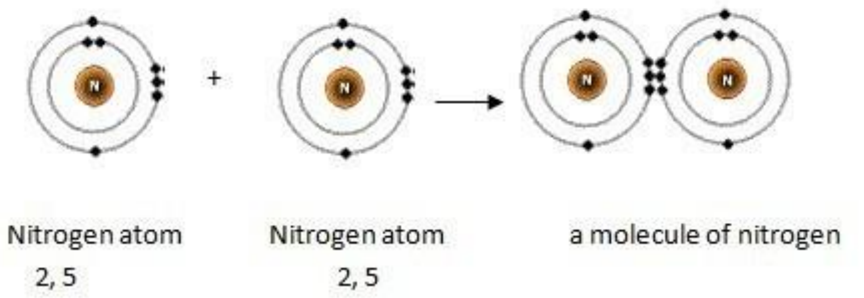
FORMATION OF N_2:

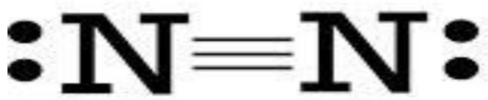
- Electronic configuration of Nitrogen is

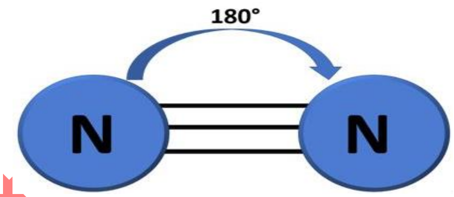
$_7N$-$1s^2\ 2s^2\ \ 2p^3$

$_7N$-$1s^2\ 2s^2\ \ 2p^3$

- Each of Nitrogen atom has 5 valence electrons.
- Thus each Nitrogen atom mutually share its three electrons with each other to complete its octet and forms N_2 molecule.







FORMATION OF NH_3(AMMONIA):

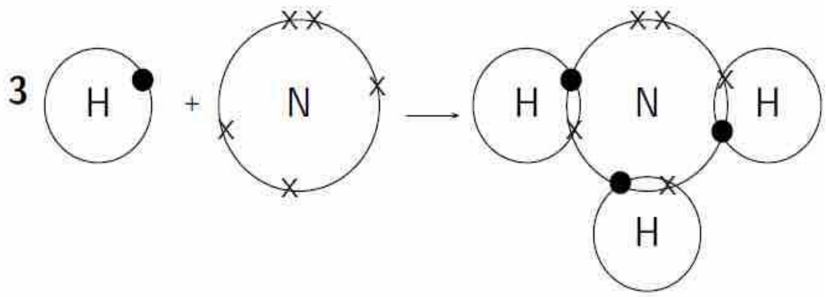
- Electronic configuration of **N** is

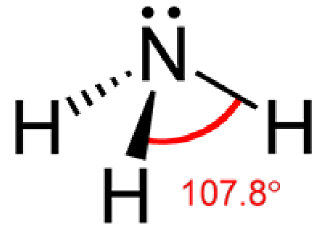
$_7N$-$1s^2\ 2s^2\ \ 2p^3$

Electronic configuration of **H** is

$_1H$-$1s^1$

- Nitrogen has 5 valence electrons and Hydrogen has 1 valence electron.
- Thus each Hydrogen atom mutually share its one electron with three electrons of Nitrogen to complete its octet and forms NH_3 molecule.





Pyramidal Shape

FORMATION OF CH4(METHANE):

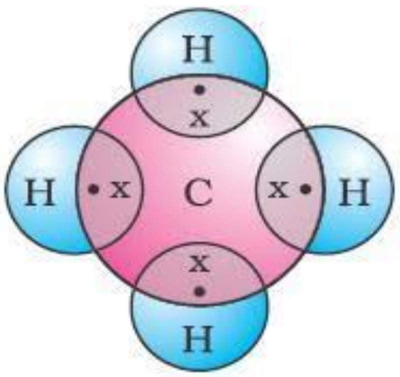
- Electronic configuration of **C** is

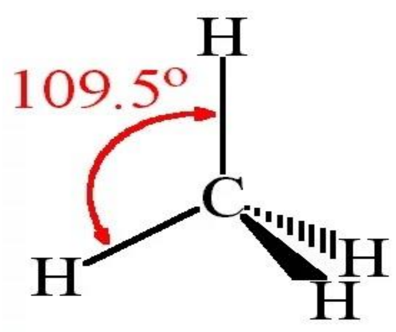
$_6C$-$1s^2$ $2s^2$ $2p^2$

- Electronic configuration of **H** is

$_1H$-$1s^1$

- Carbon has 4 valence electrons and Hydrogen has 1 valence electron.
- Thus each Hydrogen atom mutually share its one electron with 4 electrons of Carbon. to complete its octet and forms **CH_4** molecule.





(Tetrahedral structure)

IONIC BOND:

It is defined as the bond which is formed by the transfer of one or more electrons from one atom to the other atom.

Ex: NaCl KCl $MgCl_2$

FORMATION OF NaCl(SODIUM CHLORIDE)

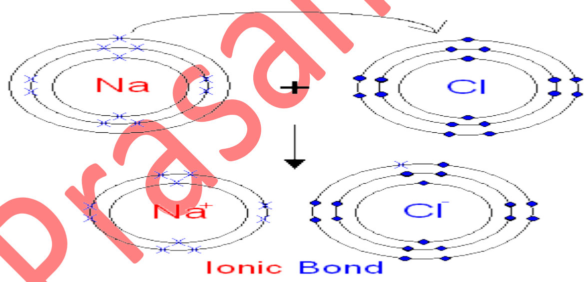
- Electronic configuration of **Na** is

$_{11}$Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$

- Electronic configuration of **Cl** is

$_{17}$Cl: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

- Sodium has 1 electron in its outermost shell.
- Chlorine has 7 electron in its outermost shell.
- **Na** donates one electron and forms **Na^+** ion(sodium cation).
- **Cl** accepts the electron donated by sodium and forms **Cl^-** ion(Chloride anion).
- Being oppositely charged **Na^+** & **Cl^-** held together by electrostatic force of attraction and forms electrovalent bond.



FORMATION OF $MgCl_2$(MAGNESIUM CHLORIDE)

- Electronic configuration of **Mg** is

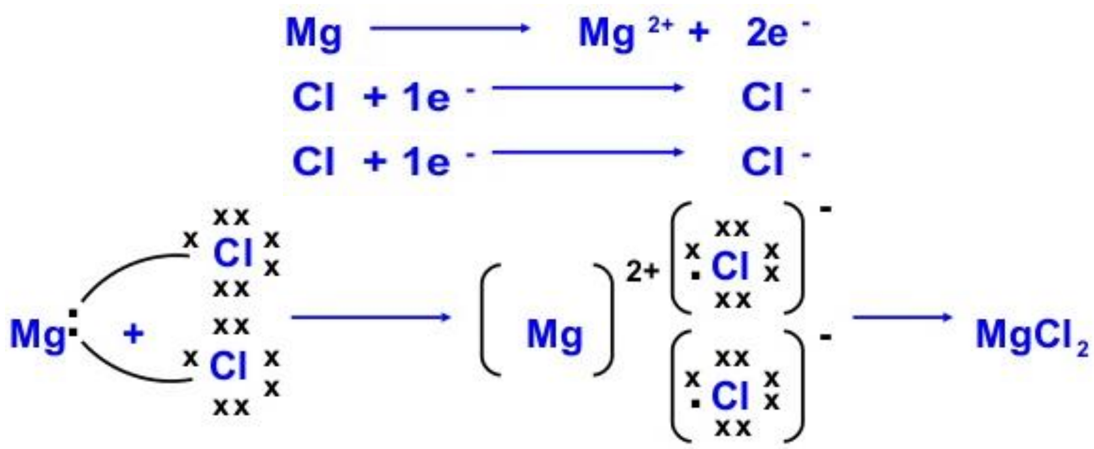
$_{12}$Na: $1s^2\ 2s^2\ 2p^6\ 3s^2$

- Electronic configuration of **Cl** is

$_{17}$Cl: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

- Mg has 2 electron in its outermost shell.
- Chlorine has 7 electron in its outermost shell.
- **Mg** donates 2 electrons and forms **Mg^{2+}** ion(sodium cation).
- Two **Cl** atoms accepts the electrons donated by Mg and forms two **Cl^-** ion(Chloride anion).

- Being oppositely charged **Mg^{2+}** & two **Cl^-** held together by electrostatic force of attraction and forms **$MgCl_2$** molecule.



COORDINATE BOND:

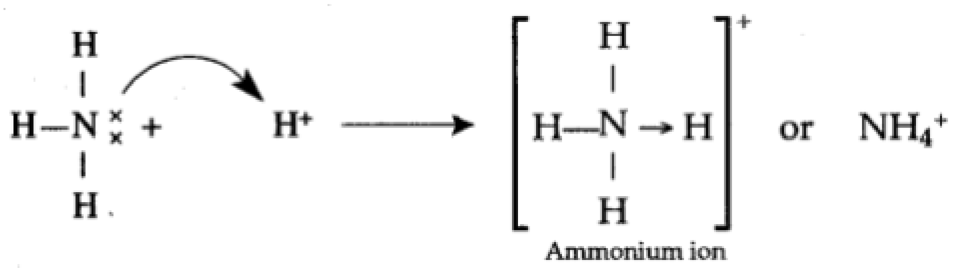
A co-ordinate bond is formed when an atom with complete octet donates its pair of electrons to the other atom. The donated pair is counted for the stability of both the atoms.

- It is formed between two dissimilar atom
- It is denoted by sign → **(single headed arrow)**
- It is directional in nature

Ex: H_2O_2 H_3O^+ NH_4^+

FORMATION OF NH_4^+ (AMMONIUM) ION

- **NH3** donates its lone pair of electron to **H^+** ion and forms **NH_4^+** ion.



FORMATION OF SULPHUR DIOXIDE(SO_2)

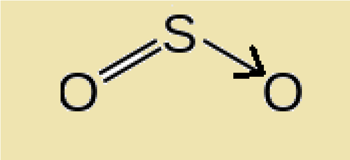
- Electronic configuration of **S** is

$$_{16}S: 1s^2\ 2s^2\ 2p^6\ 3s^2 3p^4$$

- Electronic configuration of **Cl** is

$$_{8}O: 1s^2\ 2s^2\ 2p^4$$

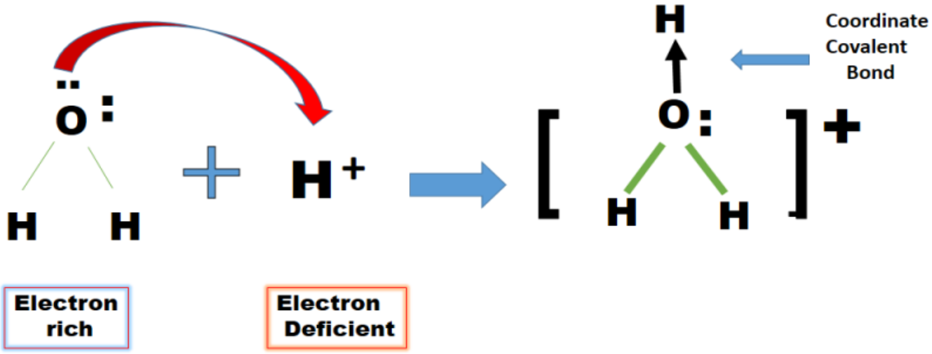
- Here Sulphur donates its lone pair to Oxygen atom



Coordinate bonding in SO_2

FORMATION OF H_3O^+ (HYDRONIUM)ION

- **H2O** donates its lone pair of electron to **H^+** ion and forms **H_3O^+** ion



CHAPTER - 3

ACID – BASE THEORIES

A. ARRHENIUS THEORY
B. LOWERY – BRONSTED THEORY
C. LEWIS THEORY.

ARRHENIUS THEORY:

According to Arrhenius theory,

1. The substances which produce **H^+ ions** (protons) in aqueous solution are called acid.

Ex: $HCl + H_2O \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

$HNO3_{(aq)} \rightarrow H^+_{(aq)} + NO_3^-{}_{(aq)}$

$CH_3COOH + H_2O \rightarrow H^+_{(aq)} + CH_3COO^-$

$H2SO4 + H_2O \rightarrow H^+_{(aq)} + SO_4^{2-}$

2. The substances which produce **OH^- ions** in aqueous solution are called base.

Ex: $NaOH + H_2O \rightarrow Na^+ + OH^-$

$KOH + H_2O \rightarrow K^+ + OH^-$

$Ca(OH)2 + H_2O \rightarrow Ca^{2+} + 2\ OH^-$

Other examples of bases are: LiOH, NH_4OH, $Mg(OH)_2$, $Al(OH)_3$, etc

3. The acids and bases which can ionize completely called strong Acids and strong Bases. Eg: - HCl, H2SO4, KOH, NaOH. The acids and bases which do not ionize completely are called weak acids and weak Bases e.g. CH3COOH, NH4OH

4. The ionisation of weak acids and bases are shown by double headed arrow.

$CH3COOH \leftrightarrow H^+ + CH3COO^-$

$NH4OH \leftrightarrow NH4^+ + OH^-$

5. According to Arrhenius theory an acid reacts with a base to form salt and water and the reaction is called **neutralization reaction.**

$NaOH_{(aq)}$ +	$HCl_{(aq)} \rightarrow$	NaCl +	H2O
base	acid	salt	water

LIMITATIONS:

1) According to this theory acid produces H^+ ion in aqueous solution but it has been found that H^+ ion cannot exist independently rather it exists in the form of Hydronium ion ($H30^+$).
2) It fails to explain the acidic and basic properties of the substances in the solvents other than water.
3) It fails to explain the acidic properties of substances like CO2, SO2 etc and basic properties of substances like NH3, CaO etc which do not contain H^+ or OH^- ion.
4) It fails to explain the neutralization reaction in the absence of water like

$NH3 + HCl \rightarrow NH4Cl$

BRONSTED - LOWRY THEORY:

According to this theory

1. Acids are the substances, which can donate a proton (**H^+**)

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

   $H_2SO_4 \rightarrow H^+ + HSO_4^-$

   $H_3O+ \rightarrow H^+ + H_2O$

   $HSO_4^- \rightarrow H+ + SO4_2^-$

2. Bases are the substances, which can accept a proton (**H^+**)

   $NH3 + H \rightarrow NH4^+$

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

   $CO3^{2-} + H^+ \rightarrow HCO^{3-}$

   $OH^- + H^+ \rightarrow H2O$

3. when an acid loses a proton, the residue formed will have a tendency to accept protons and Similarly, when a base accepts a proton residue will have a tendency to lose a proton and it acts as an Acid. This pair of acid and base which differ by a proton is called conjugate Acid Base Pair.

   Acid $\leftrightarrow$ H^+ + Conjugate Base

   Base + H^+ $\leftrightarrow$ Conjugate Acid

EX: $HCl \leftrightarrow H^+ + Cl^-$

Acid conjugate base

4. A strong acid will have a weak conjugate base and vice versa. Similarly a strong base will have weak conjugate base and vice versa.

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

Strong Acid weak conjugate base

$H_2SO_4 \rightarrow H^+ + HSO_4^-$

Strong Acid weak conjugate base

$NH3 + H^+ \rightarrow NH4^+$

Strong base weak conjugate acid

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

Strong base weak conjugate acid

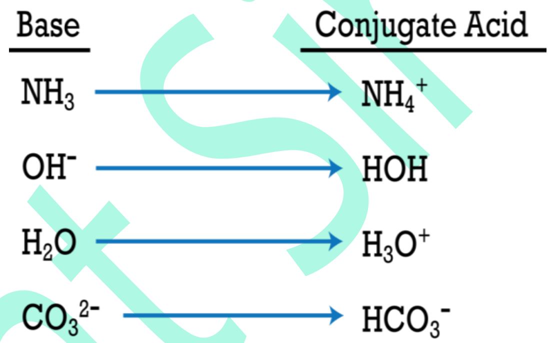
NOTE: *All Arrhenius acids are Bronsted-Lowry acids;* (HCl, HNO_3, H_2SO_4, H_3PO_4, CH_3COOH, H_2CO_3 etc.), *however the reverse is not true.*

LIMITATIONS:

1) It fails to explain the acidic nature of the substances, such as SiO_2, CO_2, SO_2, BF_3, etc. which cannot donate H^+ ion.
2) It fails to explain the basic nature of the substances, such as Na_2O, K_2O, CaO etc. which cannot accept H^+ ion.
3) It fails to explain the reaction between some acids and bases which do not give another pair of acid and base. Example: $HCl + NaOH \rightarrow NaCl + H_2O$.

BONUS POINT:

Acid	Conjugate Base
HF	F-
HBr	Br-
HNO_3	NO_3^-
$HC_2H_3O_2$	$C_2H_3O_2^-$
H_2SO_4	HSO_4^-
H_2O	OH^-



LEWIS THEORY:

According to this theory:

1. The substance which can accept a pair of electron is called acid.

Ex: All cations e.g Ca2+ , Mg2+, Al3+, Na+, H+

Molecules having multiple bond e.g. CO2, SO2 etc.

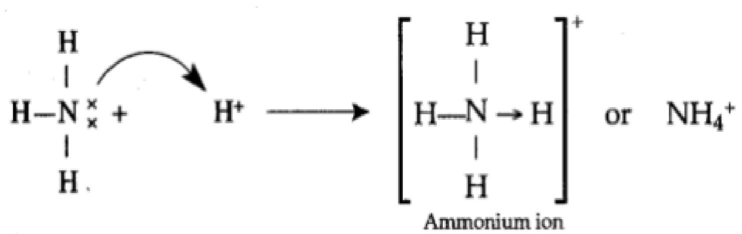
Molecules having atoms with vacant orbitals e.g. BF3, AlCl3, FeCl3, FeBr3

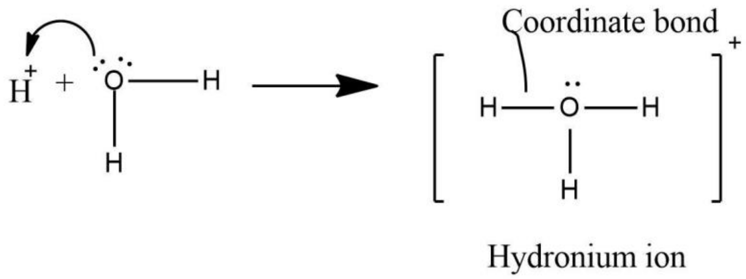
2. The substance which can donate a pair of electron is called base.

Ex: All anions e.g. Cl- OH- F- Br- , CO32- NO3- etc.

Neutral molecules having lone pair of electrons e.g. NH3, PH3,PCl3, H20 etc.

3. According to this theory, an acid reacts with a base to form a co-ordinate or dative bond.





***NOTE:** All Bronsted- Lowery bases are Lewis bases while the reverse is not always true.*

LIMITATIONS:

1. The theory fails to explain the relative strengths of different acids and bases.

2. It fails to explain reaction between some acids and bases where no coordinate bond is formed.

3. It couldn't explain the acidic nature of well-known acids like HCl, HNO_3, H_2SO_4, etc. which cannot accept electrons.

4. It fails to explain the basic nature of well-known bases like $NaOH$, KOH, etc. which cannot donate electrons.

5. It couldn't explain acid-catalyzed reactions, where H^+ ion plays important role.

NEUTRALIZATION OF ACIDS AND BASES:

According to Arrhenius Theory, acid react with bases to form salt and water. This type of reaction is called **neutralization reaction**.

Neutralization reaction may take place as follows:

1.Neutralization between a Strong Acid and a Strong Base: A strong acid reacts with a strong base to form a simple or normal salt. Its aqueous solution has a p^H of about 7 and is neutral.

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

2.Neutralization between a Strong Acid and a weak Base: A strong acid reacts with a weak base to form a acidic salt. Its aqueous solution has a $p^H < 7$ and the solution is acidic.

Example: $HCl + NH_4OH \rightarrow NH_4Cl + H_2O$

3.Neutralization between a Weak Acid and a Strong Base: A weak acid reacts with a strong base to form a basic salt. Its aqueous solution has a $p^H > 7$ and is alkaline.

Example: $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

4.Neutralization between a Weak Acid and a Weak Base: A weak acid reacts with a weak base to form a neutral salt. Its aqueous solution has a $p^H > 7$ and is alkaline.

Example: $CH_3COOH + NH_4OH \rightarrow CH_3COONH_4 + H_2O$

SALTS:

Salts are regarded as ionic compounds made up of positive and negative ions. The positive part comes from a base while negative part from an acid.

or

Salts are ionic compounds which produce cation other than H^+ and anion other than OH^- in aqueous solution.

or

Salts are the compounds formed by the neutralization reaction between acids and bases.

TYPE OF SALTS:

Salts may be classified into the following types:

NORMAL SALTS: The salt obtained by the complete replacement of all the replaceable hydrogen atoms of an acid by metal atoms is called a normal salt. These salts are obtained by the reaction between strong acids and strong bases. These salts are not hydrolyzed in aqueous solution.

Ex:

Acids	Normal salts
HCl	NaCl, KCl, $CaCl_2$, $MgCl_2$, etc
HNO_3	$NaNO_3$, KNO_3, $Ca(NO_3)_2$, $Mg(NO_3)_2$, etc.
H_2SO_4	Na_2SO_4, K_2SO_4, $CaSO_4$, $MgSO_4$, etc.
H_3PO_4	Na_3PO_4, K_3PO_4, $Ca_3(PO_4)_2$, $Mg_3(PO_4)_2$, etc.

ACIDIC SALTS: The salt obtained by the partial replacement of replaceable hydrogen atoms of an acid by metal atoms is called an acidic salt. These types of salts still contain one or more replaceable hydrogen atoms.

Ex:

Acids	Acidic salts
H_2SO_4	$NaHSO_4$, $KHSO_4$ etc.
H_3PO_4	NaH_2PO_4, KH_2PO_4, Na_2HPO_4, K_2HPO_4 etc.

Also, these are the salts obtained by the neutralization between strong acids and weak bases.
For examples: NH_4Cl, NH_4NO_3, $(NH_4)_2SO_4$, etc.

BASIC SALT: These are the salts obtained by the incomplete neutralization of poly acidic bases. Such salts contain one or more ‗OH‘ groups. Example: Ca (OH)C*l*, Mg(OH)C*l*, Zn(OH)C*l*, $Al(OH)_2Cl$ etc.
Also, these are the salts obtained by the neutralization reaction between weak acids and strong bases.
Ex: CH_3COONa, CH_3COOK, Na_2CO_3, K_2CO_3, etc.

DOUBLE SALTS: These are the molecular addition compounds obtained from two simple salts, the ions of which retain their identity in aqueous solution. Such salts give the test of all the constituent ions when dissolved in water.

Ex: Mohr's Salt [$FeSO_4.(NH_4)_2SO_4.6H_2O$], carnalite ($KCl. MgCl_2.6H_2O$), etc.

COMPLEX SALTS: These are the molecular addition compounds obtained by the combination of simple salts, the ions of which lose their identity in aqueous solution. Such salts do not give tests of all the constituent ions in aqueous solution.

Ex: $K_3[Fe(CN)_6]$, $[Cu(NH_3)_4]SO_4$, etc.

MIXED SALTS: These are the salts which give either more than one cation or more than one anion in aqueous solution.

Ex: Bleaching powder Ca (OCl)Cl, Sodium potassium sulphate $NaKSO_4$, etc.

$$Ca(OCl)Cl \rightarrow Ca^{2+} + OCl^- + Cl^-$$

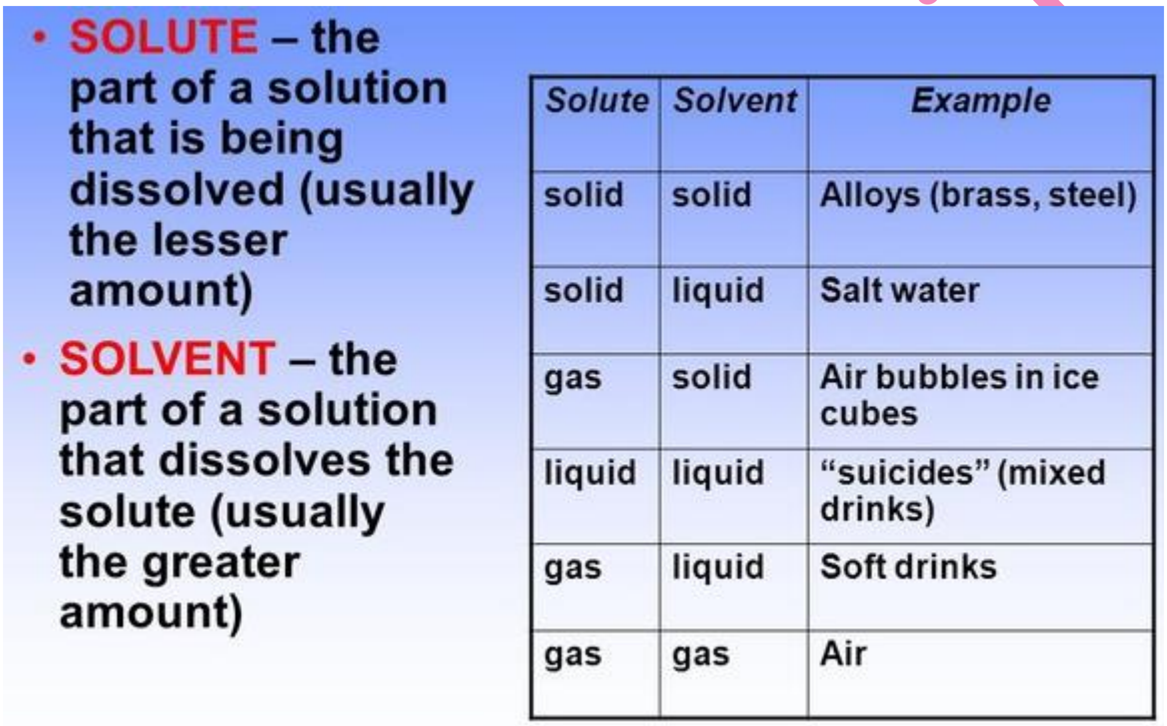
$$NaKSO_4 \rightarrow Na^+ + K^+ + SO_4^{2-}$$

CHAPTER-4
SOLUTIONS

DEFINATION

A solution is defined as the homogeneous mixture of two or more components .

In such a mixture, a **solute is** a substance dissolved in another substance, known as a **solvent**.



- **SOLUTE** – the part of a solution that is being dissolved (usually the lesser amount)
- **SOLVENT** – the part of a solution that dissolves the solute (usually the greater amount)

Solute	*Solvent*	*Example*
solid	solid	Alloys (brass, steel)
solid	liquid	Salt water
gas	solid	Air bubbles in ice cubes
liquid	liquid	"suicides" (mixed drinks)
gas	liquid	Soft drinks
gas	gas	Air

ATOMIC WEIGHT/MASS:

The atomic mass of an element may be defined as "the average relative mass of one atom of the element as compared to the mass of an atom of carbon (^{12}C) taken as 12".

Unit: amu (atomic mass unit) or simply ‗*u*'.

Element	**Symbol**	**Atomic Weight in a.m.u**
Hydrogen	*H*	*1*
Helium	*He*	*4*
Lithium	*Li*	*7*
Berylium	*Be*	*9*
Boron	*B*	*11*
Carbon	*C*	*12*

Nitrogen	*N*	*14*
Oxygen	*O*	*16*
Fluorine	*F*	*19*
Neon	*Ne*	*20*
Sodium	*Na*	*23*
Magnesium	Mg	24
Aluminium	*Al*	*27*
Silicon	*Si*	*28*
Phosphorous	*P*	*31*
Sulphur	*S*	*32*
Chlorine	*Cl*	*35.5*
Argon	*Ar*	*40*
Potassium	*K*	*39*
Calcium	*Ca*	*40*
Chromium	*Cr*	*52*
Iron	*Fe*	*56*
Copper	*Cu*	*63.5*
Bromine	*Br*	*80*
Silver	*Ag*	*108*
Lead	*Pb*	*207*

Note: When the mass is expressed in amu, it refers to the mass one atom of the element. But, when expressed in gm, it refers to the mass of 1 mole of atoms (6.023x 10^{23} atoms) of the element.

MOLECULAR MASS:

Molecular mass of a molecule is defined as the sum of atomic masses of the atoms present in that molecule.

Molecular mass of a substance is calculated by adding the atomic masses of the constituent atoms present in one molecule.

Unit: amu (atomic mass unit)

For example**:** The molecular mass of sulphuric acid (H_2SO_4) can be obtained as

H_2SO_4 = [2 × At.mass.of H] + At.mass. of sulphur +[4 ×At.mass of oxygen]

= 2 × 1+32+4×16 = **98 amu**

EQUIVALENT WEIGHT:

The equivalent weight of a substance may be defined as "the number of parts by mass of it, which combines with or displaces directly or indirectly 1.008 parts by mass of hydrogen, 8 parts by mass of oxygen or 35.5 parts by mass of chlorine."

Unit: Equivalent weight has no unit.

EQUIVALENT WEIGHTS OF ACID:

The equivalent weight of an acid is numerically equal to the molecular weight of the acid divided by the basicity.

Equivalent weights of acid = Molecular Mass / Basicity

EX:

ACID	FORMULA	MOL.Wt.	BASCITY	EQ.Wt.
Nitric Acid	HNO_3	63	1	63/1 = 63
Sulphuric Acid	H_2SO_4	98	2	98/2 = 49
Phosphoric Acid	H_3PO_4	98	3	98/3 = 32.66
Formic Acid	HCOOH	46	1	46/1=46
Acetic Acid	CH_3COOH	60	1	60/1 = 60
Oxalic Acid	COOH COOH	90	2	90/2=45
Phosphorous Acid	H_3PO_3	82	2	82/2 = 41
Boric Acid	H_3BO_3	62	1	62/1 = 62

Note: $HCOOH \rightarrow HCOO^- + H^+$

$H_3PO_3 \rightarrow 2H^+ + HPO_3^{2-}$

$CH_3COOH \rightarrow CH_3COO^- + H^+$

$H_3BO_3 \rightarrow H^+ + H_2BO_3^-$

EQUIVALENT WEIGHTS OF BASES:

The equivalent weight of a base is numerically equal to the molecular weight of the base divided by the acidity.

Equivalent weights of base = Molecular Mass / Acidity

EX:

Base	Mol.formula	Mol. wt.	acidity	Equivalent Wt.
Potassium hydroxide	KOH	56	1	56/1 = 56
Calcium hydroxide	$Ca(OH)_2$	74	2	74/2 = 37
Aluminium	$Al(OH)_3$	78	3	78/3 = 26

hydroxide

EQUIVALENT WEIGHTS OF SALTS:

The equivalent weight of a salt is numerically equal to the molecular weight of the salt divided by the total number of positive or negative charges.

Equivalent weights of salt = Molecular Mass / Total valence of cations or anions

Salt	Molecular formula	Mol. weight	Total valence of cations	Eq. weight
Sodium chloride	NaCl	58.5	1x1 = 1	58.5
Potassium carbonate	K_2CO_3	138	1x2 = 2	69
Calcium Sulphate	$CaSO_4$	136	1x2 = 2	68
Aluminium Suplhate	$Al_2(SO_4)_3$	342	2x3 = 6	57

MODES OF EXPRESSIONS OF CONCENTRATION

Concentration of a solution is the measure of the amount of solute in a given amount of solution or solvent. The concentration of a solution can be expressed in the following ways:

- **Molarity**
- **Normality**
- **Molality**

MOLARITY (M):

Molarity of a solution may be defined as "the number of gram mole of the solute present per liter of solution".

Unit = gram mole/liter or M.

Mathematically,

$$M = w * 1000 / M_s * V \quad ;$$

Where, w = weight of the solute in gram

M_s = Molecular weight of the solute.

V_{ml} = Volume of solution in ml.

NORMALITY (N):

Normality of a solution may be defined as "the number of gram equivalent of the solute present per litre of solution." It is represented by ‗N‘. Unit: - gram equivalent/liter or ‗N‘.

Mathematically,

$$N = \frac{w \times 1000}{E \times V};$$

Where, ***w = weight of solute in gm.***

V_{ml} = volume of solution in ml.

E_S = Equivalent weight of solute

MOLALITY (M):

Molality of a solution may be defined as "the number of gram mole of solute present per 1000gm (1kg) of solvent" and it is represented by the symbol ‗**m**‘. Unit: - gram mole/kg.

m= w * 1000 /M * W

Where, w = weight of solute in gm

W = weight of solvent in gm

M = Molecular weight of solute

PH OF SOLUTIONS

The ***p***p^H of a solution may be defined as -the negative logarithm of H^+ ion concentration in moles/liter or molarity.‖

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

p^H is normally used to know whether a solution acidic, alkaline or neutral in nature.

i. If $P^H < 7$; the solution is Acidic,
ii. If $P^H > 7$; the solution is Alkaline,
iii. If $P^H = 7$; the solution is Neutral.

$$[H^+][OH^-] = 10^{-14}$$

IMPORTANCE OF PH IN INDUSTRIES:

1. **In sugar Industry: -** The P^H value of the sugar cane juice should be nearly ‗7‘ i.e., it should be neural. If the P^H value of sugar cane juice becomes less than ‗7‘, the sucrose in the juice is hydrolyzed into glucose and fructose.

On the other hand, if it exceeds ‗7‘, undesirable acids and coloured substances are produced.

2. **In Paper Industries:** Paper is used in a broad array of products essential for everyday life, from newspapers, books, magazines, printing, writing papers to cardboard boxes and bags, paper napkins, sanitary tissues etc. We are daily surrounded by paper products.

   The most important use of paper is writing. The quality of paper used for printing or writing should be good and it depends on many parameters. One of the parameters is Cobb, which needs to be controlled. Cobb control is nothing but the control of quality and binding of pulp in such a fashion that whatever is written by any source such as ink, etc on paper it should not spread as well as leave its impression on back side of the paper. Cobb variation is minimized by maintaining pH of the pulp in the range of 5-6 pH. Before processing, the raw pulp has pH in the range of 7-8. This should be controlled and brought down to acidic range i.e., 5 to 6 p^H.

   Cobb control is done by addition of Alum (which is in the range of 2-3 p^H) and rosin to pulp. When alum and rosin are mixed with pulp after a certain distance pH of the mixture is measured and if it is not in the desired range the transmitter will control the Alum dosing via controller so that p^H of the pulp is maintained. Rosin on the other side has no such controlled action. It will be getting dosed to the pulp continuously in a specific quantity. It is the Alum whose dosing is controlled depending upon p^H variations.

3. **In Textile Industries:** In all textile processes in which aqueous solutions are used, balancing the pH of the solution is primary. pH control is critical for a number of reasons. The effectiveness of oxidizing and reducing agents is p^H dependent. The amount of chemicals required for a given process is directly related to the p^H. The solubility of substances, such as dyes and impurities, vary with p^H. The corrosive and scaling potential of processing solutions is also heavily influenced by p^H. All these issues affect quality and costs.

   Along with surface tension, pH plays an important role in the wetting and saturating processes. For example, caustic solutions cause interfibrillar swelling in cotton cellulose and cannot be squeezed out as easily as water, which can reduce quality in subsequent processing.

   The scouring of wool is a good example of a process where maintaining the pH value permits a better solubilization of certain impurities. For example, a p^H of 10 is considered optimum for the removal of wool wax.

   In the instance of vat dyeing, p^H controls the solubilization of the dyes. Initially, the quantity of caustic soda present must be adequate to ensure the solubility of the leuco form. Once the dye has been exhausted, the p^H is adjusted such that the dye returns to its insoluble form and is mechanically trapped in the fibre.

Between the colour kitchen and processing, controlling the p^H improves the lab-to-bulk reproducibility of colour. Monitoring and controlling pH ensures consistency of colour from batch to batch, as well.

To effectively bleach cellulose (e.g., cotton) with a minimum amount of damage, the bleaching solution must be alkaline. This keeps the hypochlorite stable and also prevents the presence of reducing groups that cause an apparently well-bleached cloth to yellow with age. Additionally, an acidic solution will form toxic and corrosive chlorine gas. Bleaching liquor is therefore usually maintained at a p^H of 9. The permanence of the white obtained is thereby increased, and the bleaching is safe. Due to environmental concerns in recent times, hydrogen peroxide bleaching has become more prevalent. Its reaction products, oxygen and water, are relatively harmless. However, hydrogen peroxide is a weak acid. Thus, its conjugate base, HO_2^-, is used to perform the actual bleaching. To ensure an adequate concentration of HO_2^-, the solution p^H must be tightly controlled. Sodium hydroxide is used to maintain the p^H at a very alkaline level of 12-12.5.

ELECROCHEMISTRY

It is the study of production of electricity from chemical energy and use of electric energy to bring about non-spontaneous Chemical transfer motions.

ELECTROLYTES

The substance which conducts electricity in its aqueous state, molten state or fused state are Called electrolytes.

"Or"

The substance which produce electricity Conducting solution when added to water is Called electrolyte.

Electrolyte is of two types :-

(1) Strong Electrolyte

The electrolytes which dissociates Completely into ions is called strong electrolyte.

Examples

1. $NaCl \rightleftharpoons Na^+ + Cl^-$
2. $CaCl_2 \rightleftharpoons Ca^{2+} + 2Cl^-$
3. $H_2SO_4 \rightleftharpoons 2H^+ + SO_4^{2-}$
4. $HCl \rightleftharpoons H^+ + Cl^-$
5. $NaOH \rightleftharpoons Na^+ + OH^-$
6. $Ca(OH)_2 \rightleftharpoons Ca^{2+} + 2OH^-$

(2) Weak Electrolyte

The electrolyte which partially dissociates into ions is called weak electrolyte.

Examples

$$CH_3COOH \rightleftharpoons CH_3COO^{-} + H^{+}$$

(Acetic Acid)

$$NH_4OH \rightleftharpoons NH_4^{+} + OH^{-}$$

(Ammonium Hydroxide)

NON-ELECTROLYTES

The substance which doesnot conduct electricity in its aqous state or fused state is known as non-electrolyte.

Examples

Sugar, Urea, Alcohol

ELECTROLYSIS

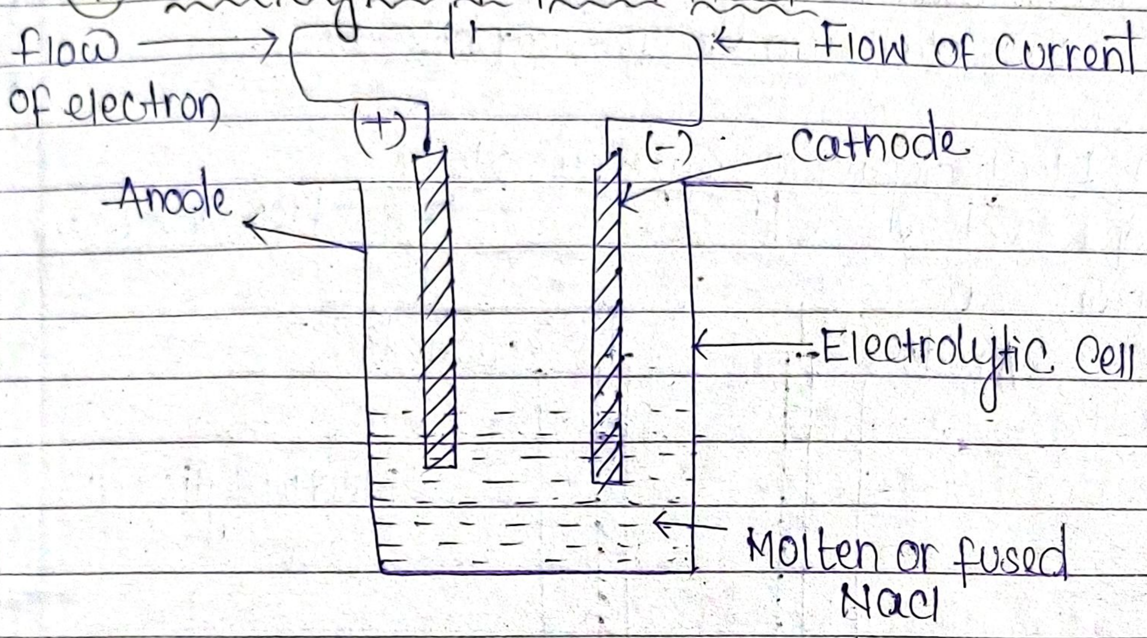
The process of chemical decomposition of electrolysis by passage of electricity is called electrolysis.

Examples

(1) fused NaCl of electrolysis

(2)

① Electrolysis of fused NaCl



In the molten state

$$NaCl \rightleftharpoons Na^{+} + Cl^{\ominus}$$

At Cathode :-

$$Na^{+} + e^{\ominus} \longrightarrow Na(s)$$

Na^{+} gains electron and deposited at cathode.

At Anode :-

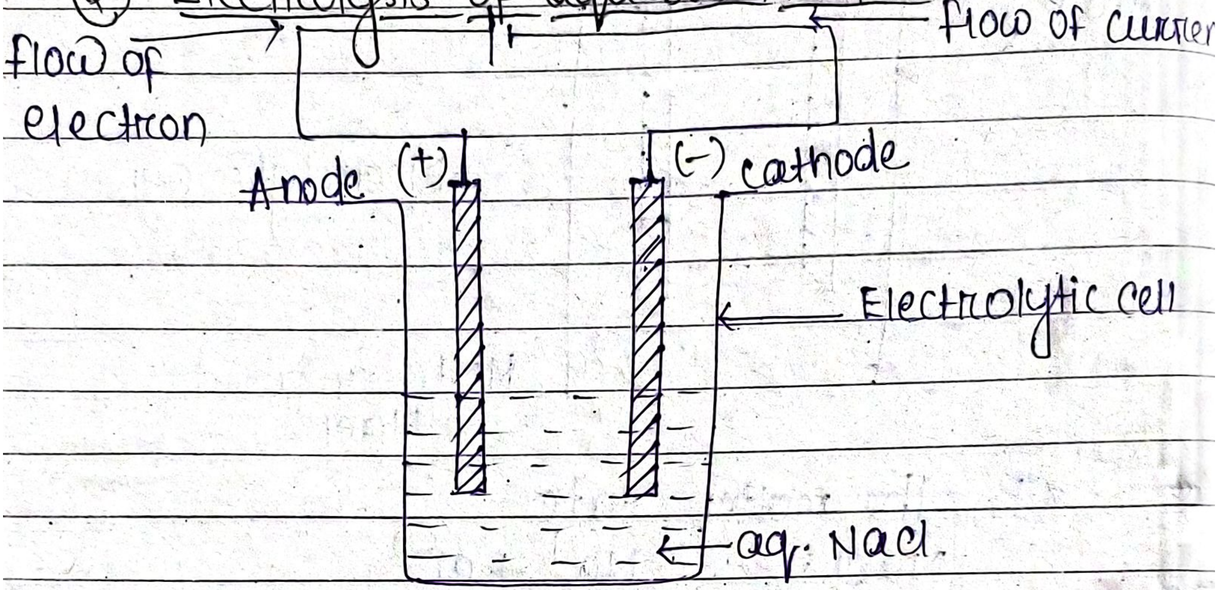
$$2Cl^{-} \longrightarrow Cl_2(g) + 2e^{\ominus}$$

Thus, in the electrolysis of fused NaCl, solution is deposited at cathode and Cl_2 gas is evolved at ~~cathode.~~ anode.

* flow of electron is always in opposited direction that of the flow of current.
* Electrodes are the Metallic rods dipped into the electrolyte solution through which electricity enters or leaves the electrolyte.

These are two types Cathode (-ve) and Anode
Ex:- platinum rod, Graphite rod.

② Electrolysis of aqueous NaCl



$NaCl \longrightarrow Na^+(aq.) + Cl^-(aq.)$

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

At Cathode

Both Na^+ & H^+ ions migrate towards Cathode

→ Since discharge ~~tubes~~ potential of H^+ is less than Na^+, thus H^+ ions are discharged.

$$2H^+ + 2e^{\ominus} \longrightarrow H_2(g)$$

At Anode

Both Cl^- and OH^- ions migrate towards anode.

→ Since discharge potential of $Cl^{\ominus}$ is less than OH^-, thus $Cl^{\ominus}$ ions are discharged.

~~$2H^+ + 2e^{\ominus} \longrightarrow H_2(g)$~~

$$2Cl^{\ominus} \longrightarrow Cl_2 + 2e^{\ominus}$$

Thus in the electrolysis of NaCl (aq.) solution, H_2 gas is evolved at cathode and Cl_2 gas is evolved at anode.

FARADAY'S FIRST LAW OF ELECTROLYSIS

The mass of the substance deposited or liberated at any electrode is directly proportional to the quantity of electricity passed.

Mathematically,

$$W \propto Q \quad \text{—— (1)}$$

Where, W = Mass of the substance deposited or liberated.

Q = quantity of electricity.

We know that, $Q = it$ —— (2)

Where, i = Current in amperes

t = Time in seconds

From equation (1) & (2)

$$W \propto it$$

or $\boxed{W = zit}$

Where z = Constant i.e. Electrochemical equivalent

→ When i = 1 ampere and t = 1 sec

Then W = z

Thus, Electrochemical equivalent of a substance is defined as the mass of the substance deposited by the passage of ampere current for 1 sec.

* → 1 faraday = 96,500 coulomb

* → 1 Coulomb or 1C = 1 A x 1 sec

FARADAY's SECOND LAW

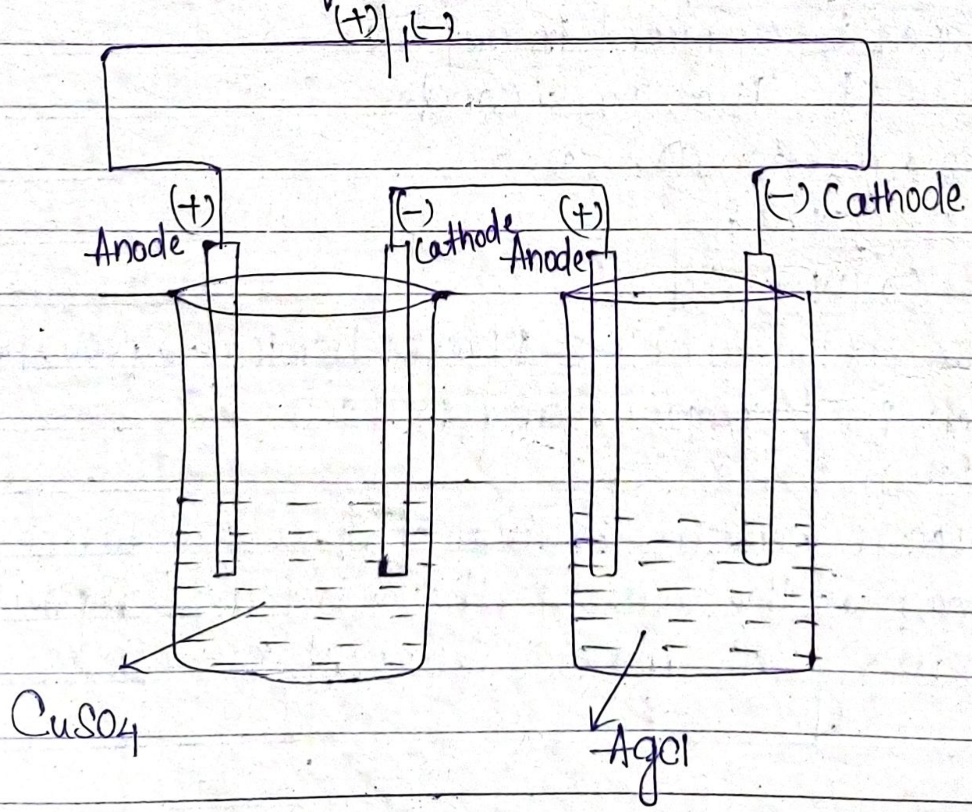
When same quantity of electricity is passed through different electrolytic solution connected in series, then the weight of different substances deposited at the different electrodes is proportional to their equivalent weight.

Mathematically,

$$W \propto E \quad \text{or} \quad \frac{W_1}{W_2} = \frac{E_1}{E_2}$$

W = Mass of the substance deposited

E = Equivalent mass of that substance



Let, two electrolytic cells containing AgCl & $CuSO_4$ connected in series.

Then, according to 2nd law of faraday

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

Where W_{Cu} = weight of Cu

W_{Ag} = weight of Ag

E_{Cu} = Equivalent weight of Cu

E_{Ag} = Equivalent weight of Ag

ELECTROPLATING

It is a process in which a thin layer of 1 metal is deposited on the surface of another metal or non-metal by appling electricity.

It is carried out generally for 3 purpose:

1. protection
2. Decoration
3. Repair

PROCEDURE

(i) The article to be electroplate is throughly cleaned with NaOH and with access of water

(ii) The Article to be electroplated is made as cathode.

(iii) The metal to be deposited on the article is made as anode.

(iv) The electrolyte should be any soluble salt of the metal to be deposited.

(v) The electrolyte is taken in an electrolytic shell or tank.

(vi) Then current is passed through the electrolyte.

CONDITION FOR BEST DEPOSITE

(i) Low current density.

(ii) Low temperature.

(iii) High metal concentration in the electrolyte.

GALVANISATION

The process of protection of Iron articles zinc coating is called galvanisation.

PROCEDURE

(i) Galvanisation process can be done by electrolysi process.

1- In this case the metal zinc to be deposited is made as anode.

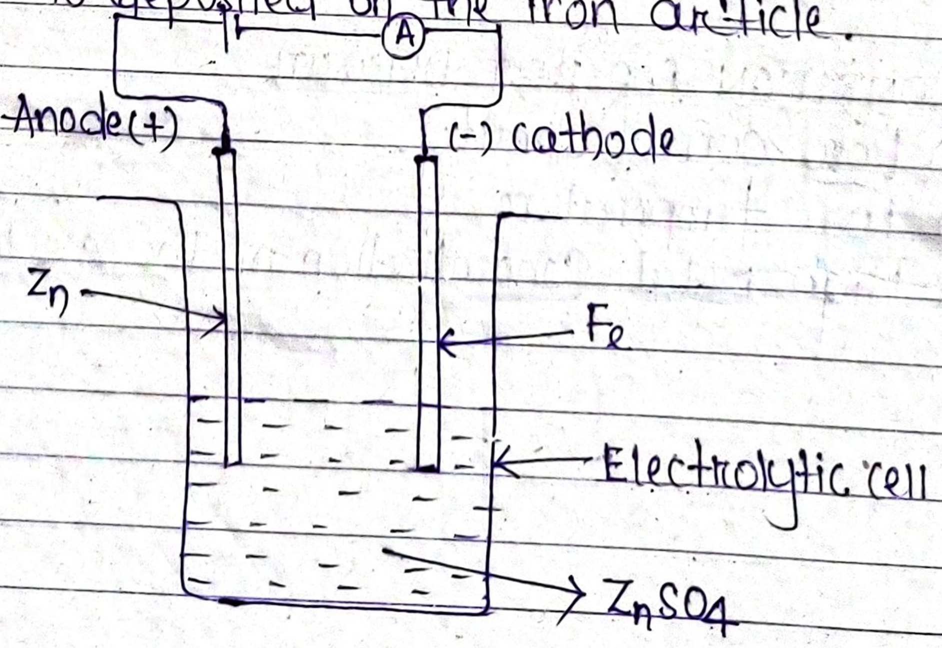
2- The iron article to be electroplated is taken as cathode.

3- The electrolyte is any soluble salt of zinc i.e. $ZnSO_4$ solution taken in an electrolytic shell.

4- The electrodes are dipped in the electrolyte solution.

5- Then current is passed through the electrolyte.

6- After some time a thin layer of zinc is deposited on the iron article.



At Anode

$Zn \longrightarrow Zn^{2+} + 2e^{\ominus}$

At cathode

$Zn^{2+} + 2e^{\ominus} \longrightarrow Zn$

$Fe_2O_3 \cdot 3H_2O$ (RUST)

$Fe \longrightarrow Fe^{2+}$

$Fe^{2+} + H_2O + \frac{1}{2} O_2 \longrightarrow Fe^{3+}$

CORROSION

It is a process which involves the convertion of metal into an undeserable compound (usually oxide), when exposed to moisture and oxygen.

Types of corrosion

(i) Atmospheric Corrosion
(ii) Waterline Corrosion

Atmospheric Corrosion

It is a process of detarioration (dg) and destruction of a material and its vital (important) properties on exposure to atmosphere.

(i) Tarnishing of silver
(ii) Development of green coating on Copper
(iii) Rusting of iron.

* Corrosion reduces the mechanical strength of the metal.

Rusting of Iron

(i) In this process Iron behaves like small electric cell in presence of water containing desolve oxygen, carbon dioxide etc.

(ii) Rusting of iron increasing in water Containing desolve oxygen.

MECHANISM

At anode,

Oxidation reaction occurs

$$Fe \longrightarrow Fe^{2+} + 2e^{\ominus}$$

At cathode,

Reduction reaction occurs

$$2H_2O + O_2 + 4e^{\ominus} \longrightarrow 4OH^-$$

(Hydroxil iron)

(i) Then the iron 2^+ ions defused OH^- ions in presence of oxygen.

(ii) The iron 2^+ is converted to iron 3^+.

- $Fe^{2+} + H_2O + O_2 \longrightarrow Fe^{3+} + 3OH^-$

* Balence — $2Fe^{2+} + H_2O + \frac{1}{2}O_2 \longrightarrow 2Fe^{3+} + 3OH^-$

- $2Fe^{3+} + 6OH^- \longrightarrow Fe_2O_3 \cdot 3H_2O$

(Rust)(Hydratade ferric oxide)

Fe^{2+} (Ferrous Iron)
Fe^{3+} (Ferric iron)

Waterline Corrosion

(i) It is Caused because of difference in oxygen Concentration.

(ii) When water is stored in a steel tank, Corrosion taken place along the line just below the level of water meniscus.

(iii) The area above the waterline is called cathode.

(iv) The area just below the waterline is called Anode.

or,

Due to these 2 electrode Corrosion takes place just below the water line that is anode.

$$Fe^{2+} + H_2O + O_2 \longrightarrow Fe^{3+} + 3OH^-$$

METHODS OF CORROSION OF METAL :-

1. ALLOYING OF METAL

Alloys can resist corrosion by two types/ways :-

(a) Homogeneity

Alloys are the homogeneous mixture or solid solution of 2 or more metals. Alloying is done with the metal which are not active to the environment.

- Ex - The rusting of iron is minimized by alloying it with chromium.

(b) oxide film

oxide film form on the surface of metal decreases Corrosion.

- Ex - Durrion : It is a silica iron alloy which is highly resistant to acid as it forms

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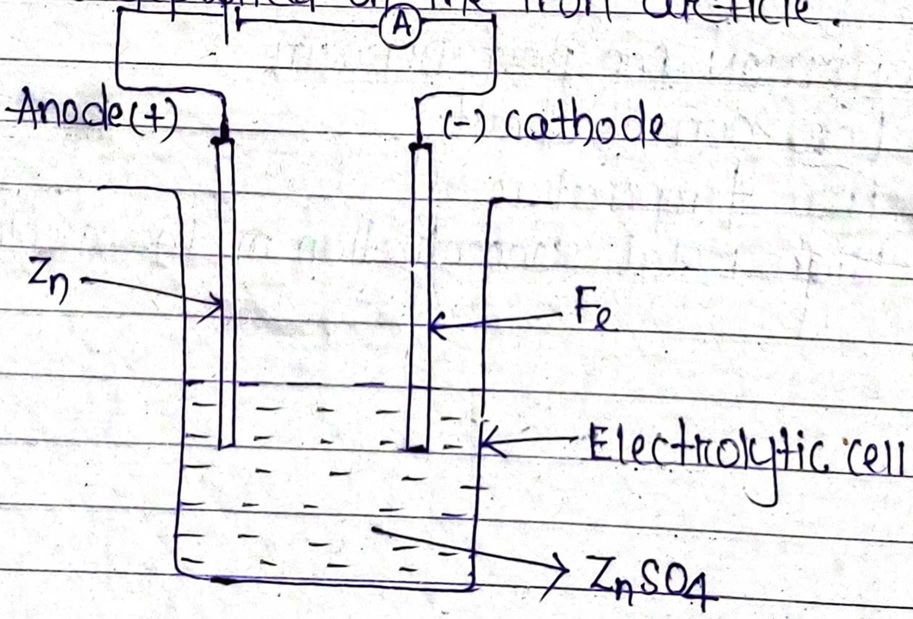

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