

~~idoblo~~

ATOMIC STRUCTURE

for 10th standard 2023-24

MATTER:- Anything having certain mass and it occupies some space, it is called as matter

* Matter is made with smallest indivisible tiny particles called as atoms

* Even though atom is smallest particle, it consists of three fundamental particles

- Electron ^{discovered by JJ Thomson}
- Proton ^{discovered by Rutherford, goldstein}
- Neutron ^{discovered by James Chadwick}

1) Electron:— (e^-) It is discovered by ^{JJ Thomson}

2) Discharge tube/ cathode ray experiment

(William Travers)

3) The charge of electron is negative

4) But the exact charge of electron is given by Neilson in oil drop experiment

charge of electron = -1.6×10^{-19} coulombs = -4.8×10^{-10} esu

Mass of electron = 1.67×10^{-27} kg (relative energy) (0.0005 amu) - negligible

The lightest particle of fundamental particles (excluding electron)

$$m(e^-) = 9.1 \times 10^{-31} \text{ gms}$$

$$= 9.1 \times 10^{-31} \text{ Kgs}$$

$$= 2 \times 10^{-23} \text{ kg}$$

$\frac{1}{1837}$ times of H molecules

$22.2 \times 10^{-23} \text{ Kgs}$

specific charge: (of e)

base charge of electron / mass of electron

If charge of an electron is -1.6×10^{-19} coulombs
mass of an electron is 9.1×10^{-31} kgs
then $\frac{-1.6 \times 10^{-19}}{9.1 \times 10^{-31}} = -1.7 \times 10^{11} \text{ C/kg}$

* Specific charge is not varies with velocity of proton speed but varies with velocity of electrons

* As the velocity of electron increases specific charge increases

2) Proton (e^+) :-

1) Discovered by Rutherford and named by Goldstein in discharge tube

2) Protons are similar to electron but opposite to direction (charge)

charge of proton :- 1.6×10^{-19} coulombs

$1.672 \times 10^{-24} \text{ gms}$
 $= 1.672 \times 10^{-27} \text{ gms-kgs}$
= Hydrogen mass

Mass of e^+ :- $1.672 \times 10^{-27} \text{ gms}$

$= 1.672 \times 10^{-27} \text{ gms-kgs}$

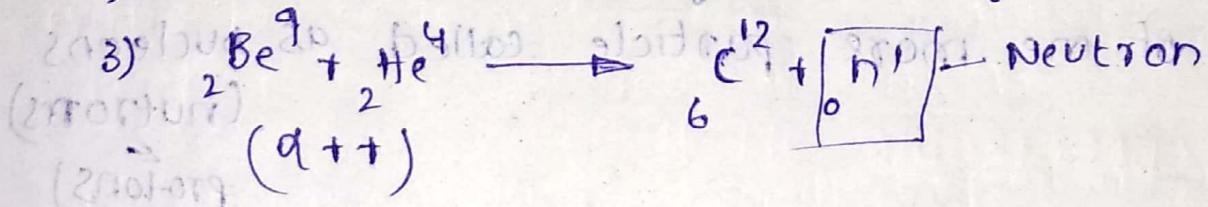
$= 1.672 \times 10^{-27} \text{ gms-kgs}$

* $m \text{ of } e^+ > m \text{ of } e^-$ (from beta decay)

so part (b) will be P mode
second option our paired 2P configuration (s)

3) Neutron: It is given by James Chadwick

2) Bombardment of 'Be' with α rays



$$(d = +ve)$$

~~2nd Be is +ve and because we have 2 protons (s)~~
~~1st He = -ve because we have 2 neutrons (s)~~

4) charge of Neutron: - zero

mass of Neutron: - $= 1.00864 \text{ amu}$

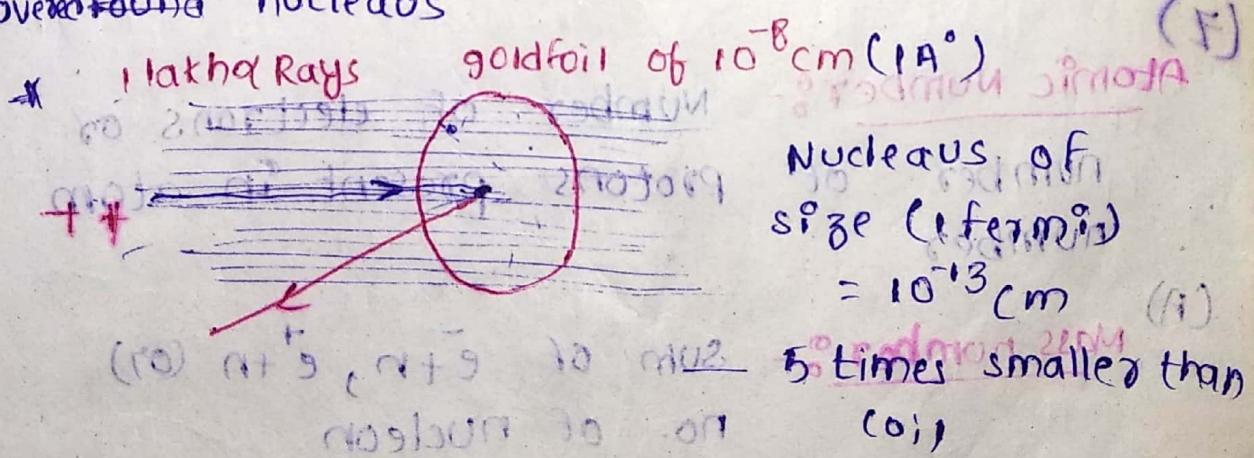
$= 1.674 \times 10^{-24} \text{ gms}$

$$= 1.674 \times 10^{-27} \text{ kgs}$$

STRUCTURE OF ATOM

\Rightarrow α Ray Scattering Experiment / Gold Foil Experiment

* This was done by Rutherford and discovered nucleus



* Even though all rays pass through coil but some of the rays deviate at centre because centre of the gold foil has a part called Nucleus (very small size)

Rutherford Atomic Model

- 1) Atom is elliptical shape
- 2) Nucleus is having +ve charge because protons are present in nucleus but nucleus is heavy because it consists of one more particle called as nucleons (neutrons & protons)
- 3) Electrons revolve around the nucleus like planets revolve around the sun

Limitations

- 1) Shape of an atom is not elliptical
- 2) Stability of an atom is not explained
- 3) formation of spectrum not explained (splitting of light)

(Z)

Atomic Number - number of protons (number)

Number of electrons or present in atom

(A)

Mass number - sum of no. of nucleon

$$A = e^+ + n^-$$

$$A = Z + n$$

$$(Z \text{ is atomic number})$$

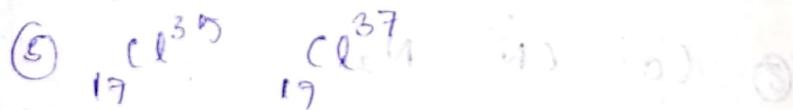
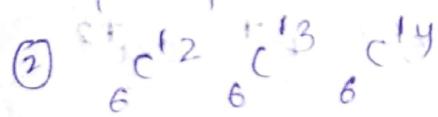
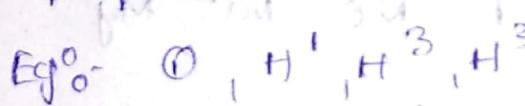
$$M = A - Z$$

$$Z X$$

ISOTOPES

-ISOTOPES:

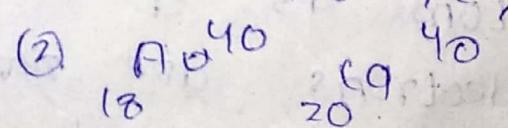
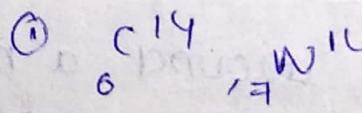
- 1) The elements which are having same atomic number but different mass numbers
- 2) These are having same no. of electrons, and protons but different no. of neutrons
- 3) Therefore these are having some chemical properties different physical properties



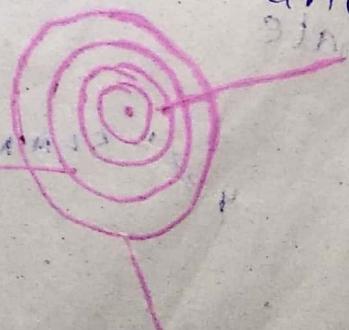
ISOBARS

- 1) The elements which are having same mass number and different atomic numbers

- 2) These are having different no. e^-, e^+, n



- 3) The elements which are having different atomic number and



ISOTONES:-

2790 TOE

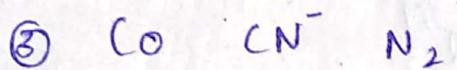
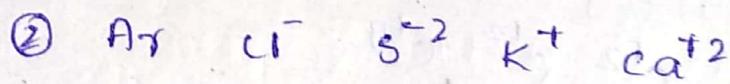
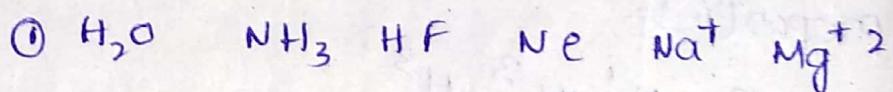
The elements which are having different atomic number, different mass number but has same no. of neutrons are called ISOTONES

EX:-	Si^{28}	P^{31}	S^{32}
	14	15	16
	16	16	16

* ISOELECTRONIC SPECIES :-

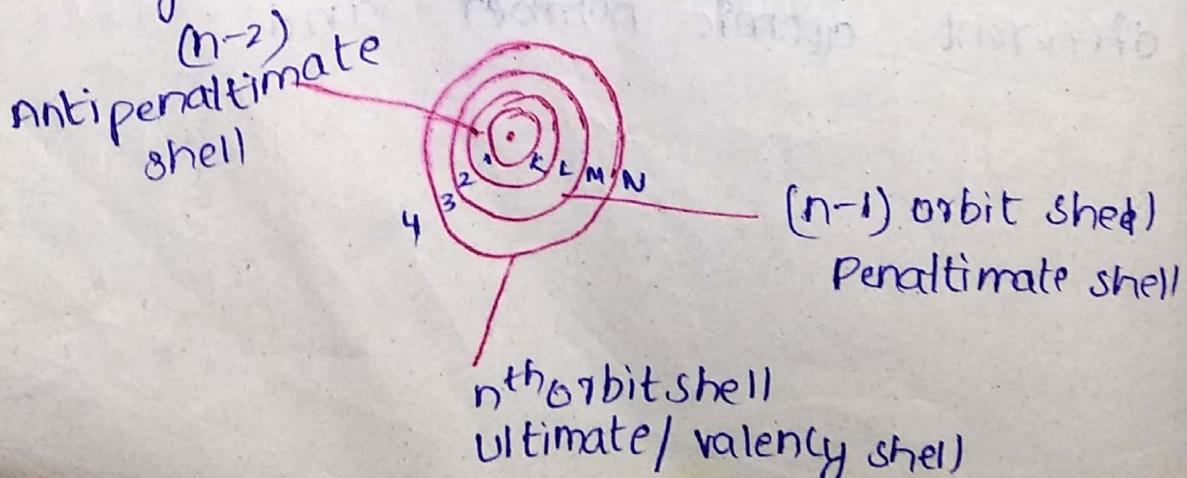
→ Species having same no. of electrons

Ex:-



BOHR'S ATOMIC MODEL

- 1) This theory depends upon Rutherford atomic model and Planck's Quantum theory
- 2) According to this theory shape of an atom is spherical
- 3) Electrons are revolving around a nucleus in a constant energy levels represented by 1, 2, 3, 4, ... (or) K, L, M, N, ...



4) when electron jumps from one orbit to other orbit, the energy of electron changes can be calculated by Plank's equation,

$$E_2 - E_1 = h\nu$$

$$\Delta E = h\frac{c}{\lambda}$$

ΔE - change in energy

c - ~~Planck's~~ velocity of light 3×10^8 m/sec

ν - frequency 3×10^{10} cm/sec

λ - wavelength

h - Planck's constant $= 6.625 \times 10^{-28}$ erg-sec

~~standard~~ $= 6.625 \times 10^{-34}$ Joule-sec

$$\boxed{\Delta E = \frac{h\nu}{\lambda}}$$

$$\text{P.S. m/s} = \frac{h\nu}{\lambda}$$

When electron is revolving around the nucleus, it possesses some angular momentum

$$mv\lambda = n\left(\frac{h}{2\pi}\right)$$

n - integer multiple (no. of orbit) - 1, 2, 3, 4

- Main energy level / Principal Quantum No.

Ex: If electron (e^-) is in ① K ② L shell then find angular momentum

① K

$$= 1\left(\frac{h}{2\pi}\right)$$

② L

$$= 2\left(\frac{h}{2\pi}\right) = \frac{h}{\pi}$$

so angular momentum increases when we move away from nucleus.

b) Radius of an orbit by using

$$r_n = \frac{n^2 h^2}{4\pi^2 m e^2}$$

$$\text{or } r_n = \frac{r_1}{n^2}$$

$$r_n = 0.529 \times n^2 A_0$$

$$r_1 = 0.529 A_0$$

$$r_2 = 0.529 \times 4 A_0$$

$$r_3 = 0.529 \times 9 A_0$$

c) The energy of n^{th} orbit

$$E_n = -\frac{2\pi^2 m e^4}{n^2 h^2}$$

$$\left[\frac{1.6 \text{ J}}{A} \right]$$

Drawbacks/ Limitations :-

1) It is applicable for only single electron atom (H, He^+, Li^{+2}, \dots)

2) This theory does not explain

Zeeman effect - splitting of light in magnetic field

Stark effect - splitting of light in electric field

3) Dual behavior of electron motion is not explained

4) Formation of fine spectrum is not given

QUANTUM NUMBERS

These are classified into four types

1) Principal Quantum Number (n) :-

- 1) This was given by Neils Bohr
- * 2) It explains size of an atom or main energy level

- 3) It is ranging from 1 to n^{th} orbit and maximum no. of electrons given by

$$2n^2 \quad (\text{ } n = \text{orbit/principal quantum no.})$$

$$K - 1 - 2n^2 = 2$$

$$L - 2 - 8e^-$$

$$M - 3 - 18e^-$$

$$N - 4 - 32e^-$$

2) Azimuthal/Angular momentum/ ^{orbital} Quantum Number ℓ :-

- 1) This was given by Sommerfeld
- 2) It explains about subshells (or) subenergy levels and shape of orbitals
- 3) It is obtained from ' n ' values

$$\ell = n - 1$$

1st orbit

$$\ell = 0$$

1s

2nd orbit

$$\ell = 1$$

2s 2p

3rd orbit

$$\ell = 2$$

3s 3p 3d

4th orbit

$$\ell = 3$$

4s 4p 4d 4f

	Max no. of electrons
s	2
p	6
d	10
f	14

- QUANTUM NUMBERS
- 4) Maximum no. of electrons in an orbital ($2(2l+1)$)
 - 5) This quantum number explains reasons for the formation of fine spectrum because each orbit is divided into no. of sub orbitals.
 - 6) Maximum no. of electrons in an orbital is 2

Magnetic Quantum Number $z - (m)$

- 1) Given by Lande
- 2) Orientation of orbitals direction
- 3) It is obtained from

$$m = -l \text{ to } l$$

- 4)

s - (l=0)	$m = 0$
p - (l=1)	$m = -1, 0, +1$
d - (l=2)	$m = -2, -1, 0, +1, +2$
f - (l=3)	$m = -3, -2, -1, 0, +1, +2, +3$

- 5) The no. of m values is obtained from $(2l+1)$ for s orbitals ($l=0; m=1$)

$$p (l=1) = 3$$

$$d (l=2) = 5$$

$$f (l=3) = 7$$

- 6) Maximum no. of orbitals / one subshell present in each orbit is given by n^2 formula.

where $n = \text{orbit no}$	No. of orbitals
1 st orbit ($n=1$)	1
2 nd orbit ($n=2$)	4
3 rd orbit ($n=3$)	9
4 th orbit ($n=4$)	16

b) This quantum no. explains about reason for formation of Zeeman effect & Stark effect.

Spin Quantum Number δ -

- 1) Given by ~~Shrod~~ Wolenbeck & Goudsmith
- 2) Orientation of electron
- 3) When two electrons are in same orbital always 1st electron is in clockwise direction
- 4) Represented by $\uparrow (\oplus) + \frac{1}{2}$
- 5) other electron is in anticlockwise direction represented by $\downarrow (\ominus) - \frac{1}{2}$

Orbitals-

Orbital	No. of nodal plane
s orbital	0
p orbital	1
d orbital	2
f orbital	3

Electronic configuration: The process of distribution (or) arrangement of electrons in a orbital is called electronic configuration.

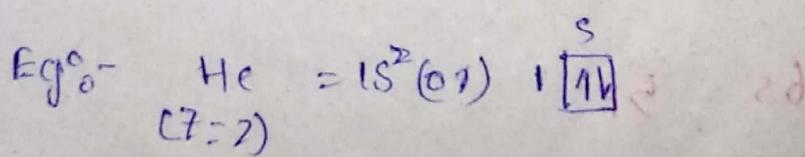
Advantages :-

- 1) It explains stability of atom.
- 2) When any atom contains only $2e^-$ or $8e^-$ in valency shell [Octet rule] it is most stable.
- 3) Any atom contains completely filled configuration like s^2, p^6, d^{10}, f^{14} are stable.
- 4) Atom contain half filled configuration like $p^3, d^5 \& f^7$ are stable.

Representation of E.C. :-

1) nl^m method

2) Box method



Rules required for E.C.

- When atoms these are divided into 3 types

Aufbau principle:-

When electrons are arranged in orbital, always free electrons enters into minimum/least energy orbital is called aufbau principle.

- The energy of an orbital is determined by using $(n+l)$ rule
- If the orbital having least $(n+l)$ value, orbital is minimum energy orbital

Ex :-

1) 2S 2P

$$(2+0)=2 \quad (2+1)=3$$

h.c.l. n.f.l.

$$2S < 2P$$

2) 3d 4s

$$n+l = 3+2$$

$$= 5$$

$$n+l = 4+2$$

$$= 6$$

$$3d < 4s \rightarrow 3d < 4s$$

⇒ If $(n+l)$ is same, consider least n value

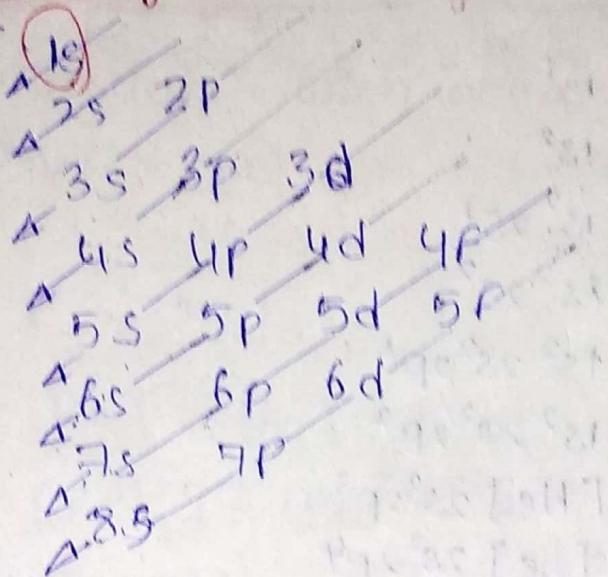
6s 5p

$$n+l = n+l$$

$$= 6 \quad = 6$$

$$5p < 6s$$

Möller diagram



Pauli's exclusive principle :-

when two electrons are in same orbital they are not having same quantum numbered values.

i.e.- when two electrons are in some orbital they have same n, l, m value but they differ in s values

[spinning of e^-]

Hund's rule:- (pairing of e^-)

The pairing of electron in a orbital when all the degenerate orbitals filled with single electrons

S - 2 e^-

P - 4 e^-

d - 6 e^-

f - 8 e^-

e.g:- Mn

Mn 1 1 1

Tin 1 1 1

Electronic configuration of 30 elements

- 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 - $[He] 2s^2 2p^3$
- 8) oxygen - $[He] 2s^2 2p^4$
- 9) Fluorine - $[He] 2s^2 2p^5$
- 10) Neon - $[He] 2s^2 2p^6$
- 11) sodium - $[Ne] 3s^1$
- 12) Magnesium - $[Ne] 3s^2$
- 13) Aluminium - $[Ne] 3s^2 3p^1$
- 14) phosphorous - $[Ne] 3s^2 3p^3$
- 15) Sulphur - $[Ne] 3s^2 3p^4$
- 16) chlorine - $[Ne] 3s^2 3p^5$
- 17) Argon - $[Ne] 3s^2 3p^6$
- 18) potassium - $[Ar] 4s^1$
- 19) calcium - $[Ar] 4s^2$
- 20) chromium - $[Ar] 4s^2 3d^4$
- 21) manganese - $[Ar] 4s^2 3d^5$
- 22) iron - $[Ar] 4s^2 3d^6$
- 23) copper - $[Ar] 4s^2 3d^9$
- 24) zinc - $[Ar] 4s^2 3d^{10}$

Application of E.C.

* E.C. decides which atom is stable

① $\text{Na} \rightarrow \text{Na}^+$

↑ ↑
11 10

$$\boxed{\text{Na}^+ > \text{Na}}$$

② $[\text{Fe}] = [\text{Ar}] 4s^2 3d^6$

$$[\text{Fe}^{+2}] = [\text{Ar}] 4s^0 3d^6$$

$$[\text{Fe}^{+2}] = [\text{Ar}] 4s^0 3d^5$$

③ $\text{Cu}(29) = [\text{Ar}] 4s^1 3d^{10}$

$$\text{Cu}^{+1} = [\text{Ar}] 4s^0 3d^{10}$$

$$\text{Cu}^{+2} = [\text{Ar}] 4s^0 3d^9$$

④ $\text{Zn}^{+2}, \text{Ca}^{+2}$

$$\text{Ca} = [Z=20] \Rightarrow \text{Ca}^{+2} = 18 [\text{Ar}]$$

$$\text{Zn} = [Z=30] \Rightarrow [\text{Ar}] 4s^2 3d^{10}$$

$$\text{Zn}^{+2} = [Z=28] \Rightarrow [\text{Ar}] 4s^0 3d^{10}$$

$$\boxed{\text{Ca}^{+2} > \text{Zn}^{+2}}$$

Formation of chemical Bond

Explained by Lewis & Kossel by
electrovalency theory

Types of chemical Bonds

They are classified into 5 types

- 1) Ionic Bond / Electrovalent Bond
- 2) Covalent Bond

Ionic Bond: Any chemical bond formed between metal and nonmetal, cation and anion, complete transfer of electron forms ionic bond

Electronegativity difference > 1.7

Ex: KCl , $NaCl$, $MgCl_2$, Al_2O_3 , Na_2O ,
 CaO etc

$$F = 4.0$$

$$O = 3.5$$

$$N = 3.0$$

$$Cl = 3.0$$

$$C = 2.2$$

$$H = 2.1$$

$$Na = 0.9$$

$$Mg = 1.0$$

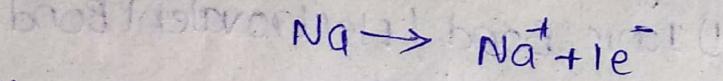
$$Ca = 0.8$$

$$Cs = 0.07$$

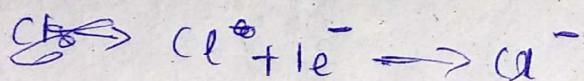
Formation of ionic compounds

1) Formation of NaCl

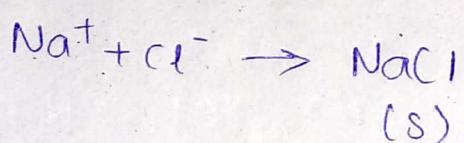
Step 1 :- Formation of cation



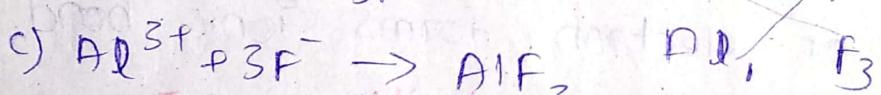
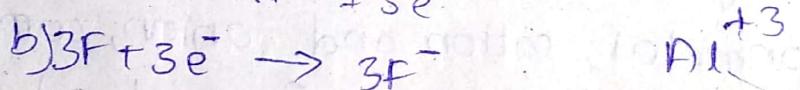
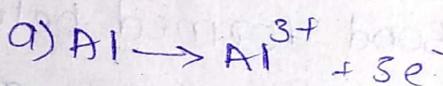
Step 2 :- Formation of anion



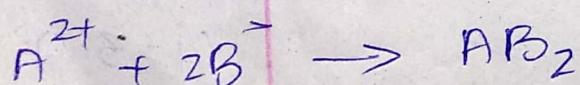
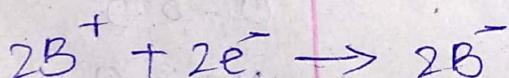
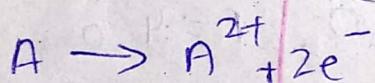
Step 3 :- Formation of neutral molecule



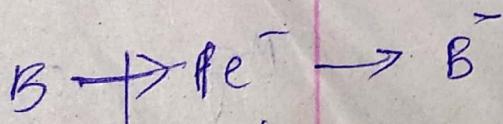
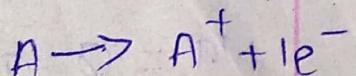
2) Al & F



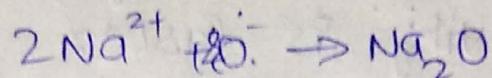
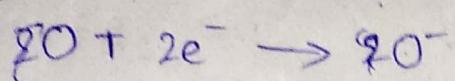
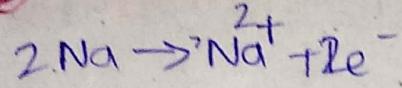
3) $(Z=12) A ; (Z=17) B$



4) $A(Z=11) ; B(Z=17)$



⑤ Na₂O

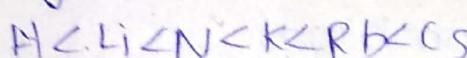


- bond dissociation

factors effecting on Ionic Bond :-

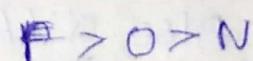
CATION

- 1) large atomic size

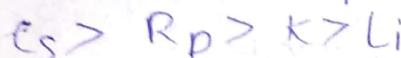


ANION

- 2) small atomic size



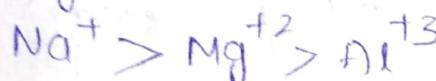
- 2) low Ionisation potential



- 2) High electron affinity



- 3) cation with small positive charge



- 3) Anion with small negative charge

Properties of Ionic compound :-

- 1) Physical state - solids

- 2) Because of strong electro static force of attraction

- 3) It is the strongest chemical bond

- 4) Ionic bond is non directional bond so Isomerism is not possible

- 5) Ionic bond soluble in polar solvents ^{water} and insoluble in non polar solvents (organic)
Benzene, Acetone, diethyl ether

- 6) Ionic compounds are bad conductors in solid state but molten state / aqueous state ^{ions} are good conductors

- 7) M.P & B.P are high
 8) Reactivity is faster

Covalent Bond :-

- 1) Same atoms combine by sharing ~~of rotation~~ of electrons and molecule formed is called homo

Ex:- O₂, N₂, H₂, S₈ (Homoatomic molecules)

- 2) Also different atoms combine together called as heteroatomic molecule

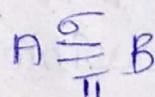
Ex:- H₂O, NH₃, CH₄, COOH

* 3) E.N ≤ 1.7 (O - 1.7) \rightarrow O-0.5 (Nonpolar)
 \rightarrow 0.5-1.7 (Polar)

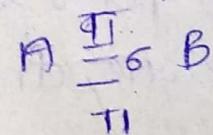
- 4) If one electron pair is shared it gives single bond



- 5) If 2 electron pair is shared it gives double bond



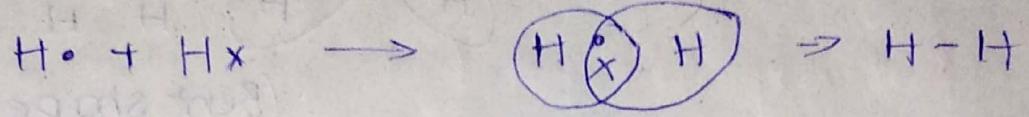
- 6) If 3 electron pair is shared it gives triple bond



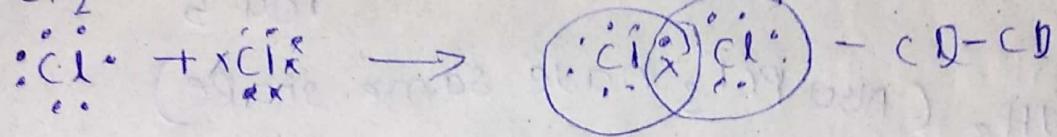
Valency: Bonding capacity

Formation of Covalent Molecule

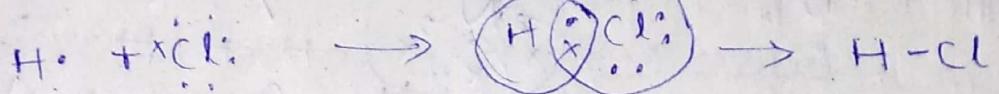
1) H_2



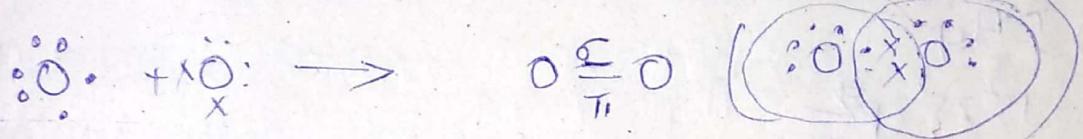
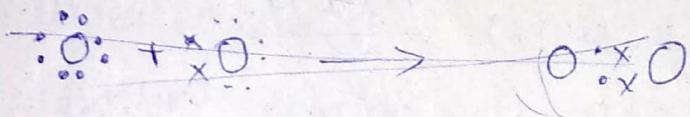
2) Cl_2



3) HCl



4) $O_2^- - 1S^2 2S^2 2P^4$

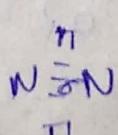
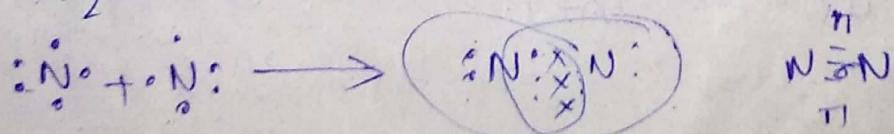


Covalent Bonds are classified into 2 types:

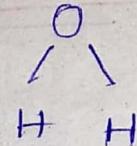
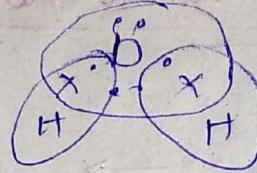
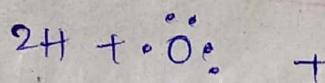
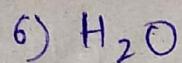
- 1) Any covalent bond formed between 2 atomic orbitals with the end to end overlapping is called σ bond
- 2) Side to side overlapping is called π bond

$$\boxed{\sigma > \pi}$$

5) N_2

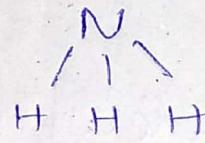
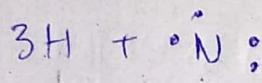
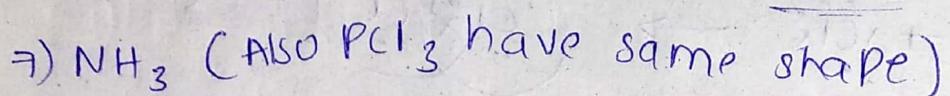


All above are linear shape and bond angle 180°



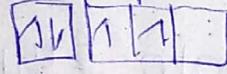
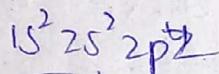
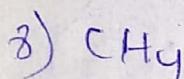
(Bent shape)

$104^\circ 5'$



(Pyramidal)

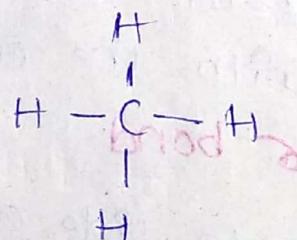
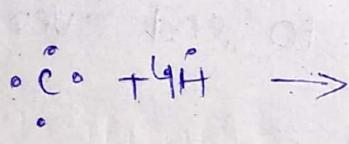
$107^\circ 4'$



- ground state



- excited state



(tetrahedral)

$104^\circ 28'$

Alkanes - $(\text{C}_2\text{H}_6) - (\text{C}_n\text{H}_{2n+2})$ - Tetrahedral

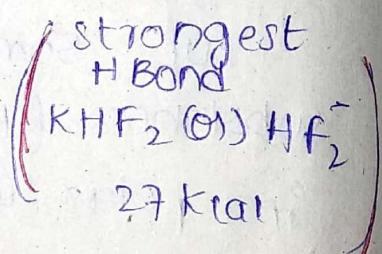
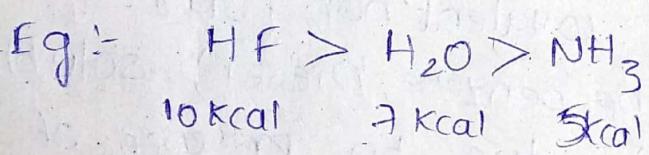
a) Ethylene molecule (C_2H_4)

Properties of covalent compounds :-

- 1) They exist in all 3 physical states
- 2) weak vanderwall's forces are present
- 3) Directionality, exhibits isomerism
- 4) covalent compounds are soluble in non polar solvents and insoluble in polar (water)
But some of covalent non polar substances (Kerosene, benzene, Diesel) soluble in polar (water) due to presence of H bonds
- 5) All covalent compounds are bad conductors of electricity except graphite
- 6) M.P & B.P of covalent compounds are less than ionic compounds
- 7) covalent reactions are slow

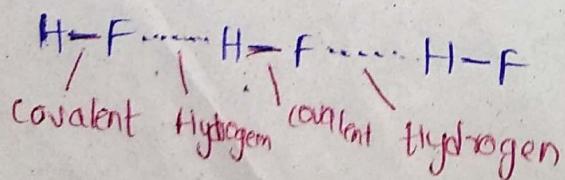
Hydrogen Bond

- Any chemical bond formed between partially positively charged hydrogen atom and most electronegative element like F, O, N is called hydrogen bond
- It is a directional bond represented by dotted line
- It is the weakest chemical bond among the all type of bond and its bond energy is 2 kcal
 - Strength of H bond depends on
 - small atomic size
 - High E.N



Classification of H Bond

Intermolecular
H Bond



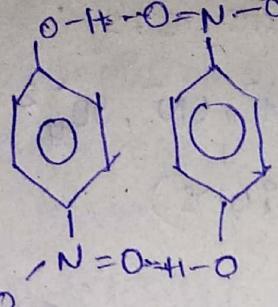
IntraMolecular
H Bond

H Bond formed
within molecules

* H Bond forms only
between molecules

Ex H_2O , NH_3 , HF_3 , Sugars

Eg:- 1) Para Nitrophenol



2) Benzaldehyde

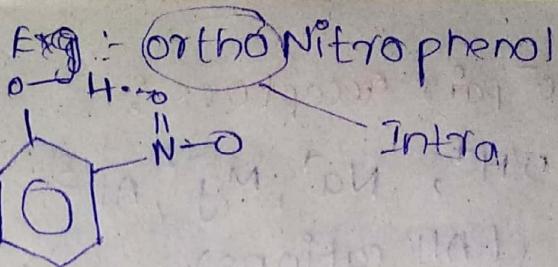
3)

Properties of H Bond :-

- 1) Ice is floated on water due to H Bond
- 2) Volume of ice is more than volume of water due to H bond due to more no. of H bonds in ice
- 3) Non polar compounds like sugar, glucose and alcohol are soluble in water due to H Bond.
- 4) Acetic acid forms dimer in benzene

Coordinate covalent Bond / Dative Bond :-

- 1) Any chemical bond formed between electron pair acceptor and electron pair donor / unequal sharing of electrons forms C.C.B
- 2) It is a directional bond (\longleftrightarrow) (directed from donor to acceptor)



2) Ortho benzaldehyde

Orthohydroxy benzaldehyde

e^- pair acceptors.

$\text{O}^{\text{H}+}$, Na^+ , Mg^{+2} , Al^{+3}

(All cations)

e^- pair acceptors.

Cl^- , OH^- , CN^- , SO_4^{-2}

(All anions)

* 2) Electrons deficient molecules

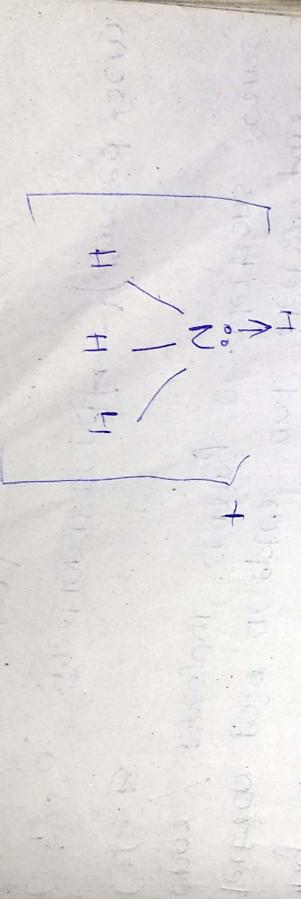
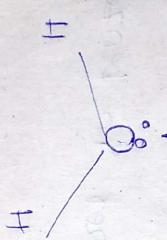
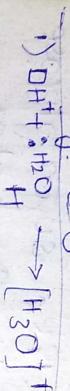
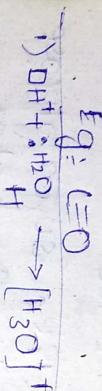
Eg: BF_3 , BCl_3 , PCl_3
 FeCl_3 , BeCl_2

3) Molecules having multiple bonds between different atoms

Eg: O_2 , N_2

2) molecules having lone pairs of electron on central atom

ROH



Metallic Bond

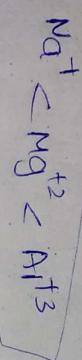
O^{6-}

- 1) Chemical Bond formed between 2 metals
- 2) Proposed by Lorenz based on electron sea model

- 3) Strength of metallic bond explained by

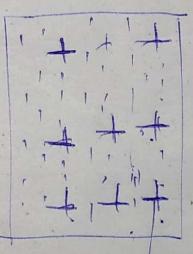
- small atomic size

- More no. of valency e^-



Electron Sea Model

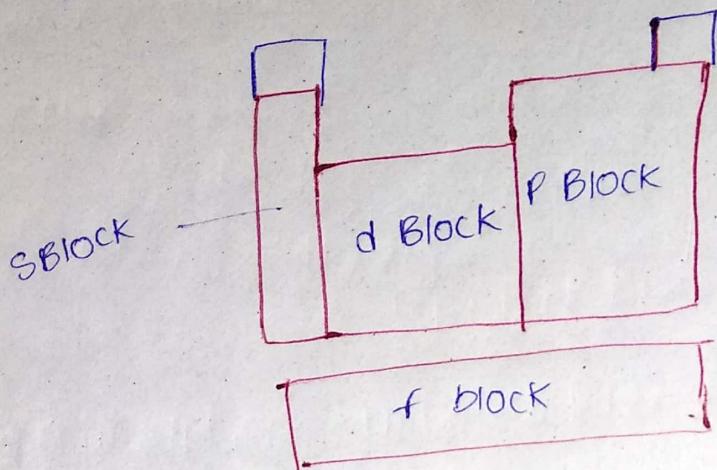
- He also explains the characteristic properties of Metals
 \rightarrow malleability, conductivity, Ductility



kernel

PERIODIC TABLE

The long form periodic table was given by Neil's Bohr based on modern periodic law i.e elements are arranged in the periodic table based on electronic configuration.



This is classified into 2 types

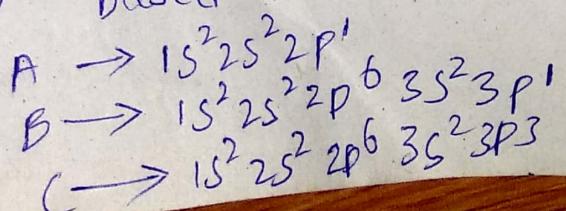
- 1) Groups (18) — 8A
— vertical columns
- 2) Periods (7)
— horizontal rows

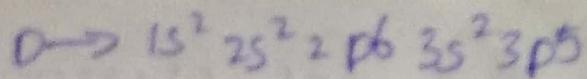
Groups characteristics:-

→ It is easier to remove electron from the last one compared to first. So electropositivity increases but electronegativity, non metallic nature, ionisation potential decreases

Atomic size, atomic radius ↑

Q) Which of the following atoms has same group based on e.c.





- Ⓐ A, B
Ⓑ C, D
Ⓒ A, C
Ⓓ B, D

2) Period characteristics :- (left to right)

short) 1st period - 2 elements - H, He

short) 2, 3, 4th period - 8 elements

long) 5, 6th, 7th period - 18 elements

long) 6th period - 32 elements

5) 7th period - incomplete

electronegativity - ↑ atomic size - ↓

Non metallic nature - ↑ atomic radius - ↓

Ionisation potential - ↑ electropositivity - ↓

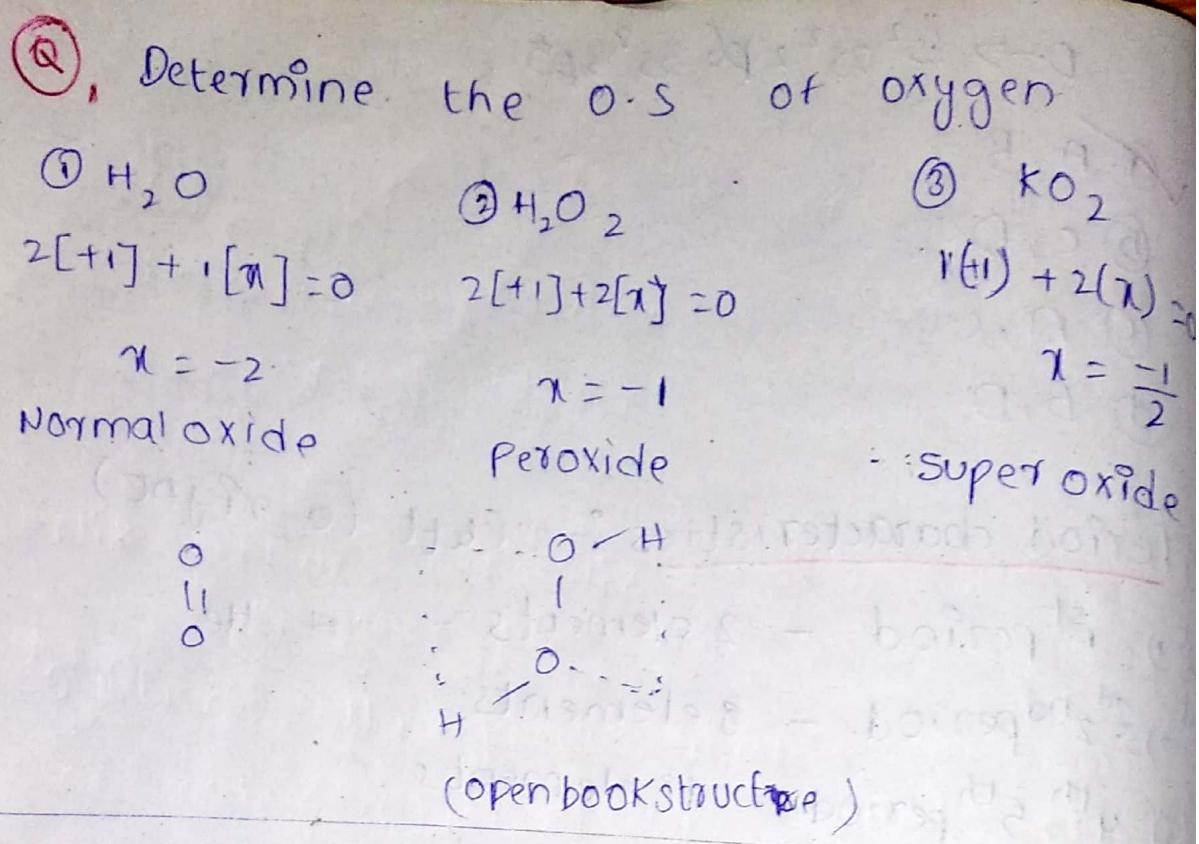
Atomic number - ↑ Metallic nature - ↓

Oxidation state

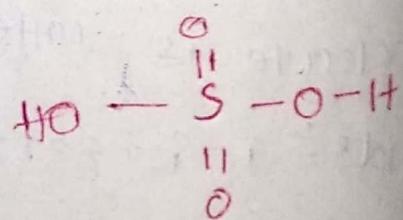
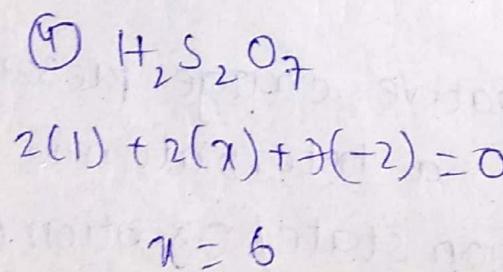
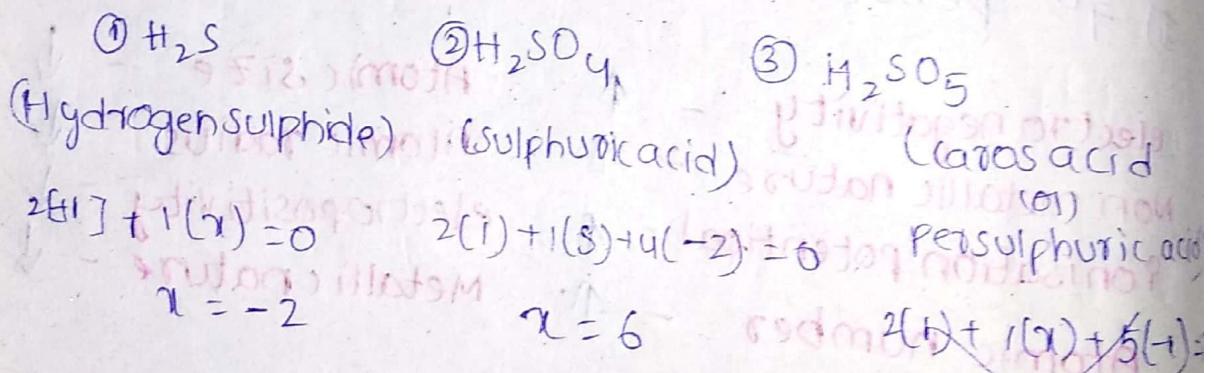
The positive or negative charge present on the central atom in a heteroatomic molecule is called oxidation state/ oxidation no.

Note: It is zero for homoatomic molecules like H₂, O₂, N₂, S₈, O₃

2) Pure metals (or) atoms Ca, Na, Mg, C

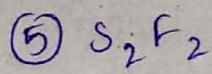


Q₂, what is the oxidation state of sulphur in



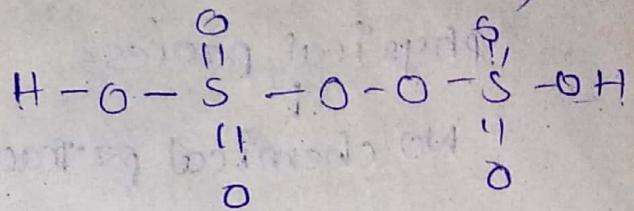
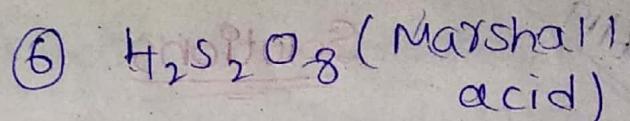
$$2(+1) + 1(-2) + 3(-2) + 2(+1) = 0$$

$$S = 6$$



$$2(x) + 2(-1) = 0$$

$$x = 1$$



$$\cancel{2[H] + 2[x] + 8[-1] = 0}$$

$$x = 3$$

$$2(+1) + 2(x) + 6(-2) + 2(-1) = 0$$

$$x = +6$$

Q3

Mn oxidation no. in $KMnO_4$ (1) MnO_4^-

Cr oxidation no. in $K_2Cr_2O_7$ (1) $(Cr_2O_7)^{-2}$

$$1(+1) + 4(-2) = -1$$

$$x = +7$$

$$2(x) + 7(-2) = -2$$

$$2x = 12$$

$$x = +6$$

Highest oxidation No. $[RuO_4] = +8$
 O_7O_4

Ruthenium tetroxide, osmium tetroxide

* some of the atoms show more than 1 valency, and such atoms are variable valency

Ex:- 1) P [PCl₃, PCl₅]

2) S [H₂S, H₂SO₄]

23/06/19

Solutions

physical process

NO chemical reactions

Ex :- Alloys, sugar in water

chemical process

chemical reactions

- Milk contains stabilizers which do not separate ghee with milk without boiling

Solution :- Homogeneous mixture of two or more components is called solution

A solution contains minimum 2 components

Solute - A quantity which is in small quantity
A quantity which undergoes physical change

Solvent -

A quantity which is in large quantity

(Q1)

A quantity which does not undergo physical change

1) Sugar + water \rightarrow Aqueous sugar solution
(solute) (solvent)

solution

2) 10ml of benzene + 100ml of kerosene
(solute) (solvent)

3) Air \rightarrow $N_2 + O_2 +$ Noble gases + (CO₂, SO₂, NO₂)
Nitrogen : ~~solute~~ solvent
O₂ + Noble gases + (CO₂, SO₂, NO₂) : ~~solute~~ solvent

Binary solution :- Any solution which contains only 2 components and such a solution is called Binary solution

Classification of Solutions :-

1) Based on nature of Solvent :-

(a) solid solution - solid (solvent)

(b) liquid solution - liquid (solvent)

(c) gaseous solution - gas (solvent)

→ solid in solid - Alloys - Brass - Cu + Zn
Bronze - Cu + Sn

Stainless steel - Fe + Cr

→ liquid in solid - Hg + Pd (Amalgam) Cu + Ni

Hygroscopic salts

($\text{CuSO}_4 \cdot 2\text{H}_2\text{O}$), gel, toothpaste

→ gas in solid - H_2 + Pt ; H_2 + Ni
(Adsorption)

→ solid in liquid - Normal Solutions

→ liquid in liquid - alcohol + H_2O , milk (fat water)

→ gas in liquid - carbonated water, aerated water

→ solid in gas - smoke

→ liquid in gas - fog (or) mist

→ gas in gas - Air

2) Based on quantity of solute

- (i) Unsaturated solution :- less quant of solute (dilute)
- (ii) saturated solution :- required quant of solut (concentrated)
- (iii) supersaturated solution :- more than required amount of solute
(trapping) (do not dissolve but settled at bottom)

******* If the temperature of solution changes, one solution is converted into other solution.

Mole concept

$$(N = 6.023 \times 10^{23} \text{ atoms/molecules})$$

Any substance contains avagardo no. of atoms/molecules (N) is called 1 mole

Ex:- 1)

$$(1 \text{ mole of O}_2 \text{ gas} = 6.023 \times 10^{23} \text{ molecules})$$

$$(2 \times 6.023 \times 10^{23} \text{ atoms})$$

$$(2 \times 32 \text{ grams})$$

∴ 1 mole of O₂ gas = 32 grams

2) calculate no. of CO₂ molecules present in 0.001 moles of CO₂ gas

$$1 \text{ mole of CO}_2 = 6.023 \times 10^{23} \text{ molecules}$$

$$0.001 \text{ mole of CO}_2 = \frac{6.023 \times 10^{23}}{0.001}$$

$$= 6.023 \times 10^{20} \text{ molecules}$$

For a given solid (or of solute)

$$\text{No. of moles} = \frac{\text{weight of solute}}{\text{gram molar mass of solute}}$$

(Example) solution contains 3.65 grams of HCl

$$\text{No. of moles} = \frac{3.65}{36.5} = 0.1 \text{ moles}$$
$$= 0.1 \times 6.023 \times 10^{23} \text{ molecules}$$

2) 0.49 grams of H_2SO_4 dissolved in H_2O

$$\text{No. of moles} = \frac{0.49}{98} = \frac{0.49}{9800} = \frac{1}{200} = 0.005 \text{ moles}$$

3) 1.06 grams of Na_2CO_3 dissolved in water

$$\text{No. of moles} = \frac{1.06}{106} = \frac{1}{100} = 0.01 \text{ moles}$$

4) 1.26 grams of oxalic acid

$$\text{No. of moles} = \frac{1.26}{126} = 1$$

(3) For a given volume of gas

No. of moles = volume of gas occupied

gram molar volume (g/m³)

for a given volume of gas no. of moles equals to

No. of moles (n) = volume of gas occupied

gram molar volume

1 mole of any gas occupies 22.4 l at 0°C and 1 atm
~~22.400 ml (cm³)~~ as gram
molar volume

An LPG cylinder contains 1200 cm^3 of
gas calculate No. of moles

$$\text{No. of moles} = \frac{11.2}{22.4} = \frac{1}{2} = 0.5$$

Concentration :- A known quantity of
solute dissolved in a fixed volume
of solution.

Molarity (M) :- No. of moles of solute
dissolved in 1 l of solution is
molarity. [formula]

Mathematically

$$M = \frac{\text{No. of moles}}{\text{litre}}$$

$$= \frac{\text{weight of solute}}{\text{Mol. wt of solute}} \times 1000$$

$$= \frac{\text{weight of solute}}{\text{Mol. wt of solute}} \times \frac{1000}{\text{Mol. wt of solute}}$$

No. of moles = $\frac{1000 \text{ g}}{\text{Mol. wt}}$

$$M = \text{moles liter}^{-1} (\text{or}) \text{ moles/litre}$$

* As temperature increases because volume
of the solution

1) For pure liquids molality equals to
 $M = \frac{10 \times \% \text{ of compound}}{\text{density(specific gravity)}}$

Excess of solute added
 on weight of substance
 addition of 10% of compound
 addition of 10% of compound x density

Eg:-

1) For 5% HCl = 36.5

$$\text{density} = 1.13 \text{ gm/cc}$$

$$M_{HCl} = \frac{10 \times 36.5 \times 1.13}{36.5} = 11.3 \text{ moles}$$

because 5/100 = 0.05 & 0.05 x 100 = 5

2) For H₂SO₄ % of H₂SO₄ = 98

$$\text{density} = 1.8 \text{ gm/cc}$$

$$M_{H_2SO_4} = \frac{10 \times 98 \times 1.8}{98} = 18$$

3) For dilute liquids Molarity is

→ the process of decreasing the concentration by the addition of pure solvent (H₂O) is called dilution process

→ The decreasing concentration is determined by

$$\boxed{M_1 V_1 = M_2 V_2}$$

$$\rightarrow \text{volume of water added} = \boxed{V_2 - V_1}$$

4) For titration process

- The process of determining concentration (molarity) of one of the solution by using end point is called Titration
- It can be determined by

$$\frac{M_1 V_1}{n_1} = \frac{M_2 V_2}{n_2}$$

(n_1, n_2) are no. of moles of reactants in a balanced chemical equation

(Ex:-) 2 moles of solute dissolved in 100 g of water find molarity

$$M = \frac{2}{100} = \frac{1}{5} = 0.2 \text{ no. of moles}$$

$$= 0.2 M$$

2) what is the molarity of 4% weight by volume of NaOH

4% (w/v) - 4 gms in 100 ml of water

$$M = \frac{4}{40} \times \frac{1000}{1000} = 1 M$$

1 M = 1 mol/l

1 mol/l solution = 1M solution

b) A centimolar H_2SO_4 solution prepared in 100 ml of water. calculate wt of H_2SO_4 required.

$$\text{Centimolar} = \frac{1}{100} \text{ M} = 0.01 \text{ M}$$

$$0.01 \text{ M} = \frac{w}{100} \times \frac{1000}{100}$$

$$w = \frac{98 \times 0.01}{100} = \frac{98}{1000} = 0.0098 \text{ gms}$$

Q) 5.85 gms of NaCl dissolved in 100ml water. 10ml of above solution is diluted to 1000ml by adding water. find molality

$$\text{Molality} = \frac{5.85}{58.5} \times \frac{1000}{10} = 10 \text{ molal}$$

$$10 \text{ molal} = 1 \text{ M}$$

$$M_1 = 10 \text{ ml}$$

$$V_2 = 1000 \text{ ml}$$

$$M_1 V_1 = M_2 V_2$$

$$100 \times 10 = 1000 \times 1$$

$$V_2 = \frac{1000}{100} = 0.01 \text{ M}$$

$M_1 > M_2$ - so diluted

$$\begin{aligned} \text{Volume of water added} &= V_2 - V_1 \\ &= 1000 - 10 \\ &= 990 \text{ ml} \end{aligned}$$

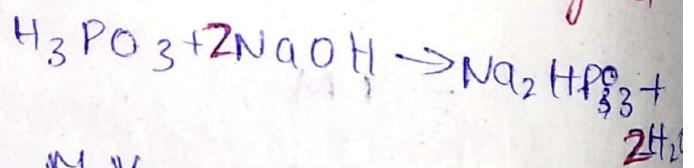
5) If 1 ml of 0.01M HCl is mixed with 1000 ml of NaCl what is resulting solution molarity

$$0.01 \times 1 = 1000 \times M_2$$

$$M_2 = \frac{0.01}{1000} = 10^{-4} M$$

* If it is reactive between two - use titration
non reactive - use dilution

6) 50 ml of H_3PO_3 is titrated with 0.1M NaOH. Volume of NaOH required to get exact end point is 25 ml. What is the molarity of H_3PO_3



$$\frac{M_1 V_1}{n_1} = \frac{M_2 V_2}{n_2}$$

$$\frac{50 \times M_1}{1} = \frac{25 \times 0.1}{2}$$

$$M_1 = \frac{25 \times 0.1}{2 \times 50}$$

$$\therefore M_1 = \frac{0.1}{4} = 0.025 M$$

② Normality = No. of gram equivalents of solute
(N) 1 l of solution

$$= \frac{\text{weight of solute} \times 1}{\text{equivalent weight of 1 l of solution}} \text{ solute}$$

$$= \frac{\text{weight of solute}}{\text{eq. wt. of solute}} \times \frac{1000}{\text{vol. of solute}}$$

$$N = \text{gram equivalent / 1 ltre}$$

* As ~~Norm~~ temperature increases, normality decreases because volume increases.

b) For pure substances

$$= \frac{10x \% \text{ of compound} \times \text{density}}{\text{Eq. wt. of substance}}$$

$$\text{Eg: } 1) \text{ For HCl } \% \text{ of HCl} = 36.5, \text{ density} = 1.13$$

$$= \frac{10 \times 36.5 \times 1.13}{36.5}$$

$$= 11.3 N$$

$(M=N)$ - because molecular weight = Eq. v. weight

Q) For H_2SO_4 , % of $\text{H}_2\text{SO}_4 = 98$, density = 1.8
 $N_{\text{H}_2\text{SO}_4} = \frac{10 \times 98 \times 1.8}{49}$

~~100 ml of 10 N sulphuric acid contains 36 N~~

Equiv. weight = mole wt / basicity / acidity

Acid - H
Base - OH

Q) For dilution & titration

$$N_1 V_1 = N_2 V_2$$

- The most convenient method for the determination of concentration is normality only (it does not depend on no. of moles)

Relation between Molarity & Normality :-

$$N \times \text{eq.wt} = M \times \text{mol.wt}$$

$$N = M \times \left(\frac{\text{mol.wt}}{\text{eq.wt}} \right)$$

* $N = M \times \text{valency}$

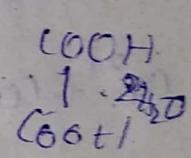
$$M = \frac{N}{\text{valency}}$$

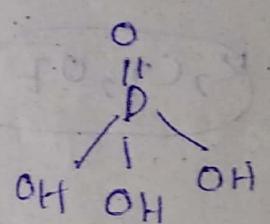
* Determination of equivalent weight

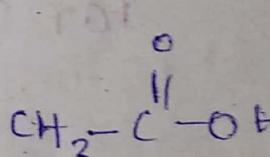
$$E = \frac{\text{Mol. wt}}{\text{Basicity (no. of replaceable } H^+ \text{ ions)}}$$

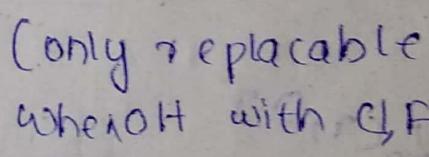
$HCl, E = \frac{36.5}{1} = 36.5 \quad (M = N)$

$H_2SO_4 \quad E = \frac{98}{2} = 49$

$H_2C_2O_4 \cdot 2H_2O \quad E = \frac{126}{2} = 63$
(oxalic acid) 

$H_3PO_4 \quad E = \frac{98}{3} = 32$
(Phosphoric acid) 

$CH_3COOH \quad E = \frac{60}{1} = 60$ 

$H_3PO_3 \quad E = \frac{82}{2} = 41$ 
(only 2 replaceable
with OH with ClF,
N, O)

$$E = \frac{\text{Mol. wt}}{\text{Acidity}}$$

$NaOH \quad E = \frac{40}{1} = 40$

$NH_4OH \quad E = \frac{35}{1} = 35$

$Ca(OH)_2 \quad E = \frac{94}{2} = 37$

(c) For Salts (Normal)

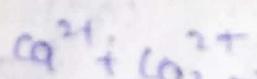
$$E = \frac{\text{Mol. wt}}{\text{total charge on cation/anion}}$$

NaCl

$$E = \frac{58.5}{1} = 58.5$$

Na2CO3

$$E = \frac{106}{2} = 53$$



$$E = \frac{100}{2} = 50$$

K2Cr2O7

for redox salts

$$E = \frac{\text{Mol. weight of salt}}{\text{charge on central atom}}$$

K2Cr2O7

$$E = \frac{294}{6} = 49$$

KMnO4

$$E = \frac{158}{7} = 22.5$$

Example

1) what is the normality of 1.3 M H_2SO_4 solution

$$N = V \times M$$

$$= 2 \times 2.3 = 4.6 N$$

2) what is the molarity of 1 N oxalic acid

$$M = \frac{N}{V} = \frac{0.1}{2} + 0.05 M$$

3) 0.63 grams of oxalic acid dissolved in 500 ml of water. what is its normality

Normality

$$\text{Normality} = \frac{\text{wt of oxalic acid}}{\text{Eq. wt}} \times \frac{1000}{500}$$

$$\cancel{0.63} \times \cancel{2} \cancel{2}$$

$$= \frac{0.63}{63} \times \frac{1000}{500}$$

$$= 0.02 N$$

4) 1.43 grams of Na_2CO_3 (286) is dissolved in 100 ml of water. If the above solution is diluted to 100 ml by addition of water. find normality

$$N_1 = \frac{w k}{E_A \cdot a t} \times \frac{1000}{V}$$

$$\frac{1.43}{2} \times \frac{100\%}{143}$$

卷之三

$$N_1 v_1 = N_2 v_2$$

$$n_2 = 0.01N$$

Volume of water added = 90 ml

5) Normality of 4% (w/v) NaOH

Amount $\frac{1000}{\text{Year}}$ Amount $\frac{1000}{\text{Year}}$

$$= \frac{cm}{h} \times \frac{1}{100}$$

二

Morality

3 11

Effect of te

Experiments

water

0.01 mg/g
100 g/m³

三

wkt

6) 50 ml of 0.1N H_2SO_4 is titrated with 0.1M of NaOH solution. what volume of NaOH solution is required to neutralise it? (10 marks)

0.1x50-0.1xV₂ 0.01mbo

Molarity

No. of moles of solute
1 kg of solvent

$$m = \frac{\text{weight of solute}}{\text{molecular weight of solute}} \times \frac{1000}{\text{wt. of solvent}}$$

Effect of temperature - not affected.

Ex- 18 grams of substance of molecular weight 40 dissolved in 500 grams of water

$$m = \frac{18}{40} \times \frac{1000}{500} = \frac{2}{5} = 0.4 m$$

2) 0.01 kg glucose was prepared in 100 gm of water

$$m = \frac{\text{wt}}{\text{m. wt}} \times \frac{1000}{100}$$

$$\text{wt} = \frac{0.01 \times 180}{1000} \times 100 = 1.8 \text{ grams}$$

Mole fraction (a) :- The ratio of no. of mol. of one component to the total no. of moles present in a solution is called mole fraction.

A binary solution prepared by using two components A & B

$$\text{no. of moles of } A = n_A$$

$$n_A = \frac{\text{no. of moles of } A}{\text{no. of moles of } A + \text{no. of moles of } B}$$

$$n_B = \frac{\text{no. of moles of } B}{\text{no. of moles of } A + \text{no. of moles of } B}$$

- No units

- Do not effect with temperature

An aqueous solution contains 36.5% of HCl. Calculate mole fraction of HCl.

$$\frac{\text{wt. of HCl}}{\text{wt. of HCl + wt. of H}_2\text{O}} = \frac{36.5}{36.5 + 63.5} = 1$$

$$n_B = \frac{\text{wt. of H}_2\text{O}}{\text{m.wt of H}_2\text{O}} = \frac{63.5}{18} = 3.51$$

$$\eta = \frac{n_A}{n_A + n_B} = \frac{1.5}{4.5} = \frac{2}{5} = 0.4$$

a) $\frac{18 \times 2}{18 + 2}$

$$\eta_{H_2O} = 1 - \eta_{HCl}$$

$$= 1 - 0.8 = 0.2$$

9.5 grams of H_2SO_4 dissolved in 18 grams of water calculate mole fraction of H_2SO_4 & water.

$$Molar mass of H_2SO_4 = 98 g/mol$$

$$Molar mass of water = 18 g/mol$$

Mole fraction of H_2SO_4 = $\frac{9.5}{9.5 + 18} = 0.35$
Mole fraction of water = $\frac{18}{9.5 + 18} = 0.65$

$$\eta = 0.35$$

b) Molar mass of H_2SO_4 = 98 g/mol
Molar mass of water = 18 g/mol

$$\eta = \frac{9.5}{9.5 + 18} = 0.35$$

$$[A]$$

9.5 gms of H_2SO_4 dissolved in 18 gms of water

Molar mass of H_2SO_4 = 98 g/mol
Molar mass of water = 18 g/mol

ACIDS & BASES

Modern theories of acids & bases

Arenius theory

- Any substance which gives H^+ ions in aqueous solution is called **ACID**.
 - Any substance which gives OH^- ions in aqueous solution is called **BASE**.
- Ex:- NaOH , KOH , NH_4OH - (Alkalies)
 $\text{Ca}(\text{OH})_2$, $\text{Mg}(\text{OH})_2$, $\text{Mg}(\text{OH})_2$

(Alkali - water soluble bases)

Strength of acids & bases

- Consider HA (general acid) It undergoes the following dissociation



- Dissociation constant $[K_a]$ = $\frac{\text{Product of product}}{\text{Product of reactants}}$ or

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

Strong acid \rightarrow If $K_a = 100\%$

e.g:- HCl , H_2SO_4 , HNO_3 , HClO_4

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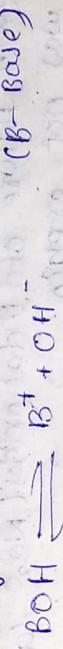
~~weak acid~~

Tb $K_a \approx 100(0) < 100\%$

$\text{Fe}^{2+}, \text{H}_3\text{PO}_4, \text{KCN}, \text{CH}_3\text{COOH}$, oxalic acid.

$$\text{H}_2\text{O}_3, \text{H}_3\text{BO}_3 \text{ etc.}$$

similarity



$$K_b = \frac{[B^*][OH^-]}{[BOH]}$$

Strong base

$K_s = 1$ (or) 100% (strong base)

$\text{Eq} \frac{1}{2} \text{ NaOH}, \text{ KOH}$

$K_b < 1 \times 10^{-5}$: weak base

$$\text{Eq}^{\circ} \text{ } \text{NH}_4\text{OH}, \text{Mg}(\text{OH})_2, \text{Al(OH)}_3,$$

Zur Theorie der

Neutralisation

10 According to this theory, an acid

and base combines to get give salt such as process is called

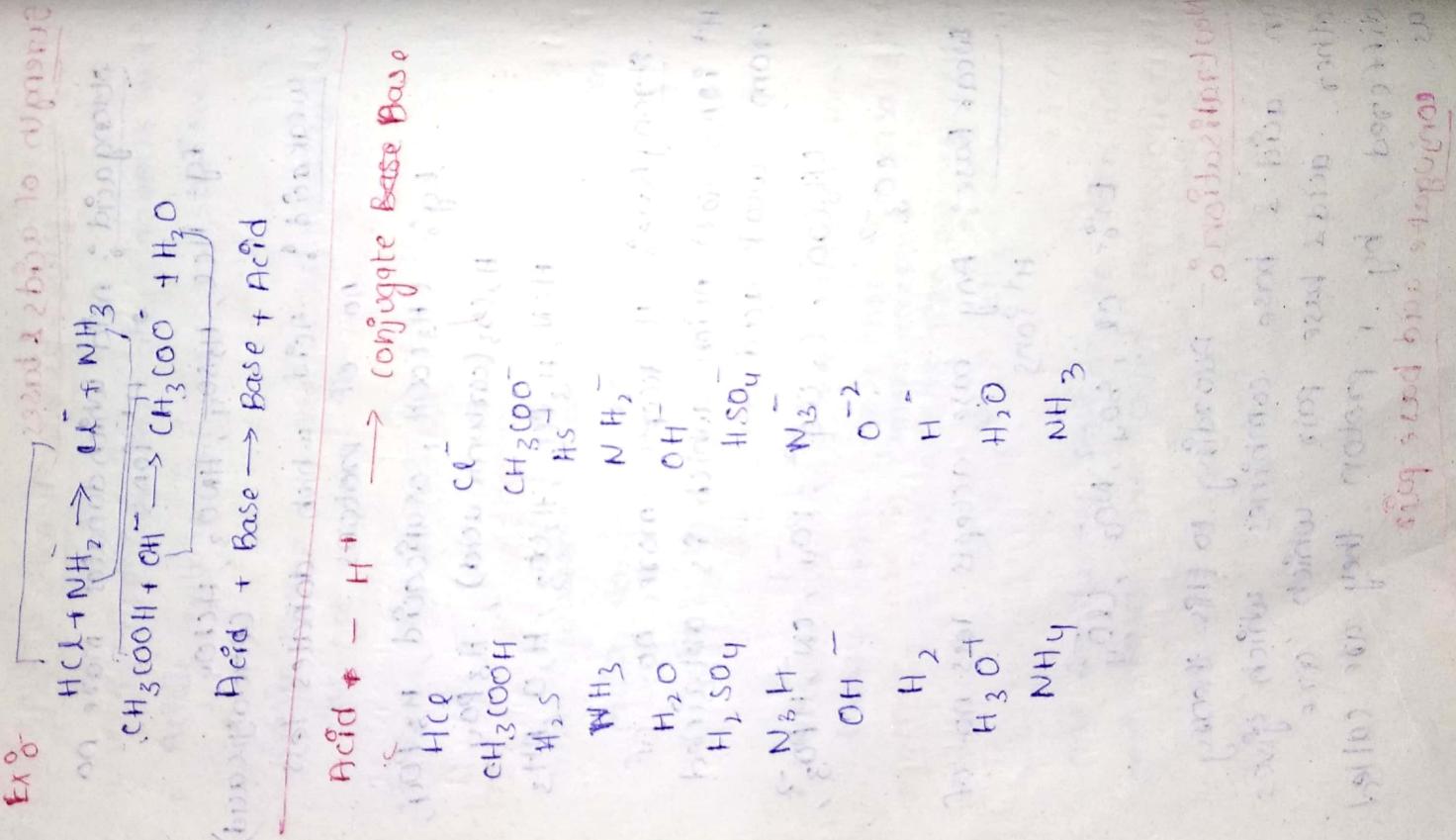
and water, see

Neutralisation process

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

$$\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COONa} + \text{H}_2\text{O}$$

(sodium acetate)



Conjugate acid + conjugate base \rightarrow water

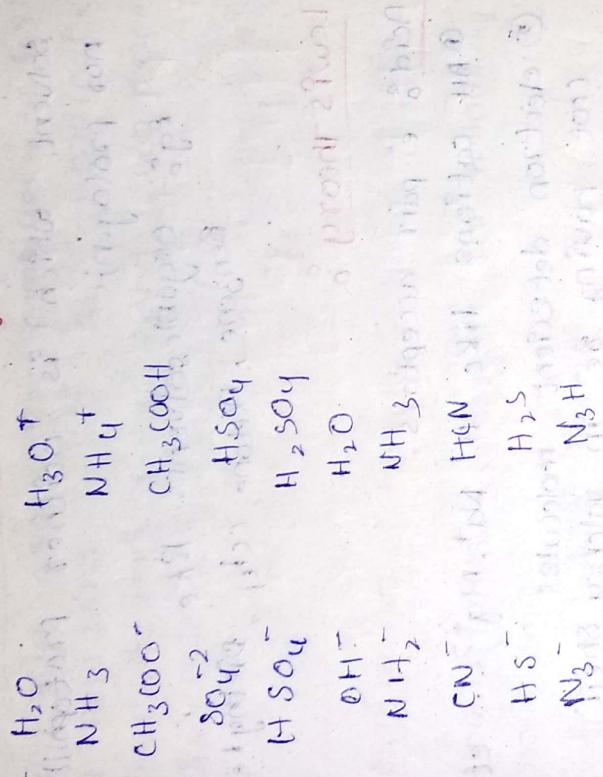
Acid + base \rightarrow salt + water

Acid + OH⁻ \rightarrow salt + water

Base + H⁺ \rightarrow salt + water

Conjugate acid + conjugate base \rightarrow water

Base + H⁺ → Conjugate Base + H₂O



(Solvent)

According to this theory, they are divided into 4 types of solvents.

1) Prophyllic solvent O^- - Proton accepting solvent

Ex:- H₂O, NH₃

2) Protogenic solvent O^- - Proton generating solvent

Ex:- All acids except oxalic acid
like HCl, H₂SO₄, Cl₂COH, Dichloro(ROH)₂
such as pyridine, Dimethyl formamide, H₂O, NH₃

3) Amphotropic solvent O^- - Solvent which is protophyllic and protogenic.
Ex:- H₂O, NH₃

④ Nprotic solvent or Non polar solvent

Solvent which is neither protophilic nor protonic

Eg:- Organic solvents like

Benzene, Acetone, $\text{C}_6\text{H}_5\text{Cl}$, Diethyl ether

Lewis theory O^-

Acid O^- e^- pair acceptors

① All cations like H^+ , Na^+ , Mg^{2+} , etc

② Electron deficient molecules not having $8e^-$ in valency shell

Eg:- BF_3 , BCl_3 , AlCl_3 , FeCl_3 , Molecules having multiple bonds between atoms

Eg:- CO_2 , SO_2 , SO_3 & CO (Non metal oxides)

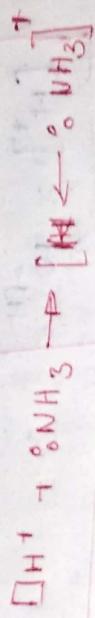
Base O^- e^- pair donor molecules

① All anions like OH^- , CN^- , Cl^- , etc

② Molecules having lone pairs of electrons on central atom like water, ammonia, etc. Eg:- $(\text{ROH})_3$, H_2O , NH_3 , $(\text{NaO})_3$

③ Multiple bonds between same atoms
Eg:- O_2 , N_2 , C_2O_4 , $\text{C}\equiv\text{C}$

Neutralisation: According to this theory an acid and base combines to form a new bond called as co-ordinate covalent bond/dative bond, such process is called neutralisation.



Determination of pH value

Consider water undergoes the following dissociation



$$\text{Dissociation constant } (K) = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$$

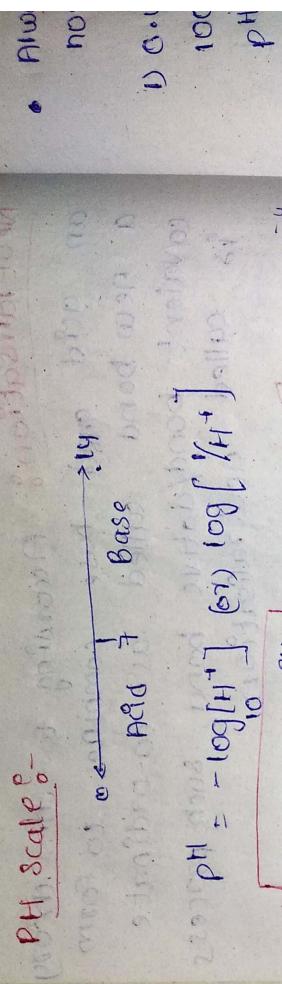
$$K[\text{H}_2\text{O}] = [\text{H}^+][\text{OH}^-] \rightarrow 0$$

$$\text{At } 25^\circ \text{C} \quad K_w = 10^{-14} \text{ (mole/lit)}^2$$

$$A \cdot 10^{-14} = [\text{H}^+][\text{OH}^-]$$

$$\begin{aligned} -\log(10^{-14}) &= -\log[\text{H}^+][\text{OH}^-] \\ \text{pH} \log 10 &= -\log[\text{H}^+] + -\log[\text{OH}^-] \\ \text{pH} &= \text{pH} - \text{pOH} \end{aligned}$$

pH scale



$$\text{pH} = -\log[\text{H}^+]$$

$$(e) \log \left[\frac{1}{\text{H}^+} \right] = \frac{-\text{pH}}{10}$$

$$10^{-4} = \frac{10^{-\text{pH}}}{10^2}$$

If pH value decreases by 4 to 2
then $[\text{H}^+]$ concentration

- (a) \uparrow_2 $[\text{H}^+] = 10^{-4} = 0.001$
- (b) $\downarrow 100$ $[\text{H}^+] = 10^{-2} = 0.01$
- (c) ~~$\uparrow 100$~~ \Downarrow_2 pH

$$\text{pOH} = -\log_{10}[\text{OH}^-] \quad (e) \log \left[\frac{1}{\text{OH}^-} \right]$$

- * If $[\text{H}^+]$ ion concentration is increased from 10^{-4} to 10^{-2} then its pH value is decreased by 2^{th} times

3) pH

1000

- * If $[\text{H}^+]$ ion concentration is decreased from 10^{-2} to 10^{-3} then its pH value increased by 1 unit

- * If pH is increased from 4 to 9,
then $[\text{H}^+]$ ion concentration decreased by 1000 times

76

- Always pH is determined by using only normality

1) 0.49 grams of H_2SO_4 dissolved in 1000 ml of water, what is the pH of resultant solution

$$N = \frac{0.49}{98.09} \times \frac{1000}{1000}$$

$$[H^+] = 0.01N$$

$$\text{Now } pH = -\log [10^{-2}] \\ = 2$$

2) pH of 0.0005M H_2SO_4 solution is

$$[H^+] = 0.0005 \times 2 \\ = 0.001$$

$$pH = -\log [0.001] \\ = 3$$

3) pH of 0.01M NaOH solution

$$N = 2$$

$$pH = [OH^-] = 10^{-2}$$

$$pH = 12$$

$$n_6 1 \text{ mole is taken} \\ [OH^-] = 10^{-14} \quad pH = 0$$

9) pH of 0.05M Ba(OH)_2 solution

$$\begin{aligned} [\text{H}^+] &= 0.05 \times 2 \\ &= 0.1 \\ \text{pH} &= -\log [10^{-1}] \end{aligned}$$

$$\text{pH} = 13$$

5) what is the pH of resulting solution when 1ml of 0.1M HCl + 1000 ml of NaOH

$$N_1V_1 = N_2V_2$$

$$\begin{aligned} N_2 &= \frac{1 \times 0.1}{1000} \\ &= 10^{-4} \end{aligned}$$

$$\text{pH} = 4$$

~~$$\text{pH} = 4$$~~

6) what is the pH of 0.01M CH_3COOH

$$\boxed{\begin{aligned} [\text{H}^+] &= \sqrt{K_a \cdot C} && \text{weak acid} \\ [\text{OH}^-] &= \sqrt{K_b \cdot C} && \text{weak base} \\ C &= \text{concentration} \end{aligned}}$$

$$[\text{H}^+] = \sqrt{10^{-4} \times 0.01} = \sqrt{10^{-6}} = 10^{-3} \text{ M}$$

$$[\text{H}^+] = c \cdot q \quad (\text{degree of dissociation})$$

$$[\text{OH}^-] = c \cdot q$$

Buffer solution - Any solution which maintains constant pH value is they are classified into 2 types

Acidic Buffer solution - $\text{pH} < 7$

- composition: weak acid + its conjugate salt



→ Basic Buffer solution or $\text{pH} > 7$

- composition - weak base + conjugate salt



→ It is used in complexometric titration

pH of Buffer solution can be known by

Henderson's equation

$$\text{pH} = \text{pK}_a + \log \frac{[\text{salt}]}{[\text{acid}]}$$

$$\text{pH} [\text{salt}] = [\text{acid}]$$

$$\text{pH} = \text{pK}_a \quad (\text{dissociation constant})$$

$$= -\log \text{K}_a$$

$$pH = pK_b = -\log K_b [OH^-]$$

(Q) which one of the pK_a values of acetic acid gives high concentration

a) 10.76

b) 3.74 ~~concentration \downarrow~~ (concentration \uparrow)
c) 5.73 ~~concentration \downarrow~~ (concentration \uparrow)

d) 4.76 ~~concentration \downarrow~~ (concentration \uparrow)

$K_{a1} = 1.3 \times 10^{-5}$

poison in water

f) $H_2O + CO_2 \rightleftharpoons H_2CO_3$

Water dissociates \rightarrow H^+ and OH^-

Water dissociates

at 25°C $\approx 10^{-7}$

$K_w = 10^{-14}$

Logarithm of K_w is called pK_w

$pK_w = 14$

Water is a weak acid

$[H^+] = [OH^-] = 10^{-7}$

CORROSION

- corrosion is a reaction due to environment over the metal surface either through chemically or through electrochemically damaging the metal
- In corrosion metal reacts with non metals like $\text{O}_2, \text{H}_2\text{O}, \text{N}_2, \text{Cl}_2$
- Noble metals are the metals which don't corrode
Eg: Gold, silver, platinum
- sometimes silver also gets corroded by H_2S (Formation of Ag_2S)
- Reactions
 - $\text{Mg} + \text{N}_2 \rightarrow \text{Mg}_2\text{N}$ (Formation of Mg₃N₂)
 - $\text{Mg} + \text{O} \rightarrow \text{MgO}$ (Formation of MgO)
 - $\text{Al} + \text{Cl} \rightarrow \text{Al}_2\text{Cl}_3$ (Formation of Al₂Cl₃)
 - $\text{Zn} + \text{R}_2\text{O}_2 \rightarrow \text{ZnO}_2$ (Formation of ZnO₂)
- corrosion is a spontaneous process for attaining stability
- Metal + environment \rightarrow metallic compound + energy
- so corrosion is an exoergic process, exothermic
- During corrosion metal loses electrons, so it undergoes oxidation
- corrosion is reverse to extraction of metals

* In minerals, ~~not~~ ~~not~~ ~~not~~ all ones are
minerals of sub oxides or even
metals.

* In extraction, all characteristics are get
reversed compared to corrosion

Classification of corrosion

- * Metal reacts with any gas and undergoes corrosion in the absence of conducting medium is called Dry corrosion.
- * Dry corrosion with oxygen is called Oxidation corrosion.

Metal Oxide

Protective		Non Protective	
stable	unstable	porous	volatile
Al_2O_3 , TiO_2 , SnO_2	Fe_2O_3 , Pb_2O_3 , Pt	acidic	Leaves
Fe_2O_3		basic	Moisture
		alkaline	
			Molybdate
			Polybarium

Wet corrosion (electrochemical corrosion)

- two separate areas form during F.C. corrosion

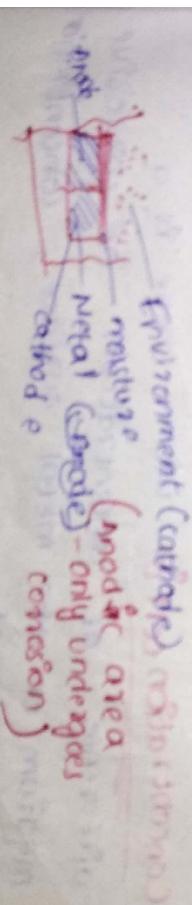
i.e. - anodic & cathodic area

- In anodic area, metal atoms loose electrons to form positively charged metal ions M^+
- i.e. oxidation takes place at anodic area
- Redox reactions will migrate from anodic area to cathode through metal
- At cathodic area electrons are attracted by oxygen containing substance to give anions. i.e. reduction takes place X^-



- positive ions from anodic area and negative ions from cathodic area join together to form corrosion product (By electrostatic/ coloumbic forces)

REDOX REACTION - CORROSION



During wet cell electrochemical cells forms on metal surface

There are 3 different ways in which new electrochemical cell formation can take place

i) concentration cell

ii) stress cell

iii) composition cell

Composition cell when 2 different & are connected and exposed to corrosive medium one metal begins to act as cathode and other as anode

- According to electrochemical series anode & cathode formation takes place
- The corrosion in composition cell also termed as galvanic cell corrosion

- As the electro potential difference increases between anode & cathode, corrosion rate increases

Zn & Cu - slower

Zn & Ag → faster

Aluminum & Copper

EDTA

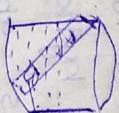
Concentration cell when there is a difference of concentration of corrosive medium on either metal surface, concentration cell is formed

- Area under more concentration becomes anode and less concentration becomes cathode

• Concentration / oxygen cell - differential aeration corrosion

Pust

C. R. I. gets corroded first



gets corroded

STRESS CELL

The formation of anodic and cathodic areas on the metal surface in presence of stress



pitting corrosion

(Pinholes)

• more accelerated and pits formed due to localized corrosion

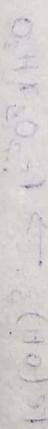
e.g. Aluminium spoon kept in NaCl
(pickles)

Intergranular corrosion

• not exposed to surface
e.g. usually exhibited in wetting welded stainless steel

• corrosion only happens inside the grain boundary

2) due to oxygen difference between air and water



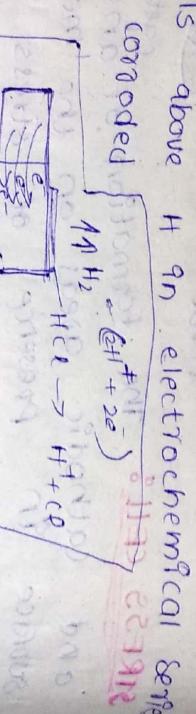
(dissolution).

• Good bonding of paper to paper will
increase life of paper - 200 years compared to
100 years without good bonding

Rusting of iron

- Any metal which undergoes corrosion in air then hydrogen gas is released

Metals above H in electrochemical series gets corroded



- For rusting of iron moisture & oxygen is required

Iron rusting is catalysed by both moisture and oxygen.



\downarrow



- If terrous hydroxide is formed which is the rust formed



(yellow rust)

- If supply of oxygen is limited, black anhydrous magnetite is formed (mixture of ferrous oxide & ferric oxide) (Fe_3O_4) - black rust

Factors Affecting Corrosion

Metallic factors

Metallic factors

- 1) Nature & Electropotential of c.
- 2) Resistivity $\propto \frac{1}{c}$
- 3) Surface finish $\propto \frac{1}{c}$

4) Relative of cathodic areas

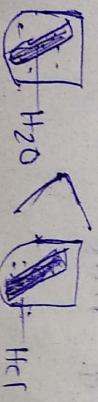


- 5) Nature of corrosion product
- 6) Passivity :- offering corrosion resistance by a metal
- 7) If a metal undergoes less corrosive than expected

• Al is more passive in O_2 compared with, due to stable, protective metallic oxide

ENVIRONMENTAL FACTORS

- 1) Moisture
- 2) Impurities
- 3) P_H of corrosive medium
- 4) Temperature
- 5) Ability of corrosive medium to dissolve corrosion product



more corrosive

METHODS OF PREVENTION OF CORROSION

Using metals of high purity :-

High pure metals are less corrosive than impure metals

Alloying :-

Chromium and nickel are alloyed with iron to get stainless steel which has high corrosive resistance

Dec

3) Modification of Environment

Decarburization :- Exclusion of oxygen at cathode

Deactivation :- by adjustment of temp by means of addition of chemicals which are capable of combining with

Ex :- Hydrazine, sodium sulphate

(2) Non metal
(Inorganic)

4) Dehumidification :- Reduction of moisture content of air

Ex :- Al_2O_3 , silica gel, ammonium

(Organic)

5) Alkaline Neutralization :- Prevents the corrosion by neutralizing the acidic character of corrosive environment
Ex :- Ammonia, sodium hydroxide and lime

6) Use of Inhibitor

Protect
Isolate
environment

(1) Metallic

Ques No. 8

- Q) Use of Inhibitors - By spraying inhibtors a thin layer of film form on surface which prevents from corrosive environment
- b) Protective coatings - coated surface isolates metal from corroding environment

① Metallic coatings

Anodic coatings - Zn on Fe (Galvanizing)

Cathodic coatings - (Sn on Fe) - If coating is continuous

method
mechanical
agitation
which

- with O_2
- ② Non metallic coatings -
- a) Inorganic coatings -
- Phosphate coatings - Zn, Fe, Steel
 - chromate - Zn, Al, Mg
- Oxide coating - Steel & copper
- modified coatings - Al, Zn, Mg by chromate coatings

③ Organic coatings -

Paints

⑤ Cathode Protection

a) Sacrificial Anodic Protection

More active metal is coated on base metal which acts as anode and gets corroded.

Eg:- Mg, Zn, Al are sacrificial anodes.

b) Impressed Current Cathodic Protection

By using cathode, base metal made ~~acts~~ as cathode which prevents corrosion of base metal.

