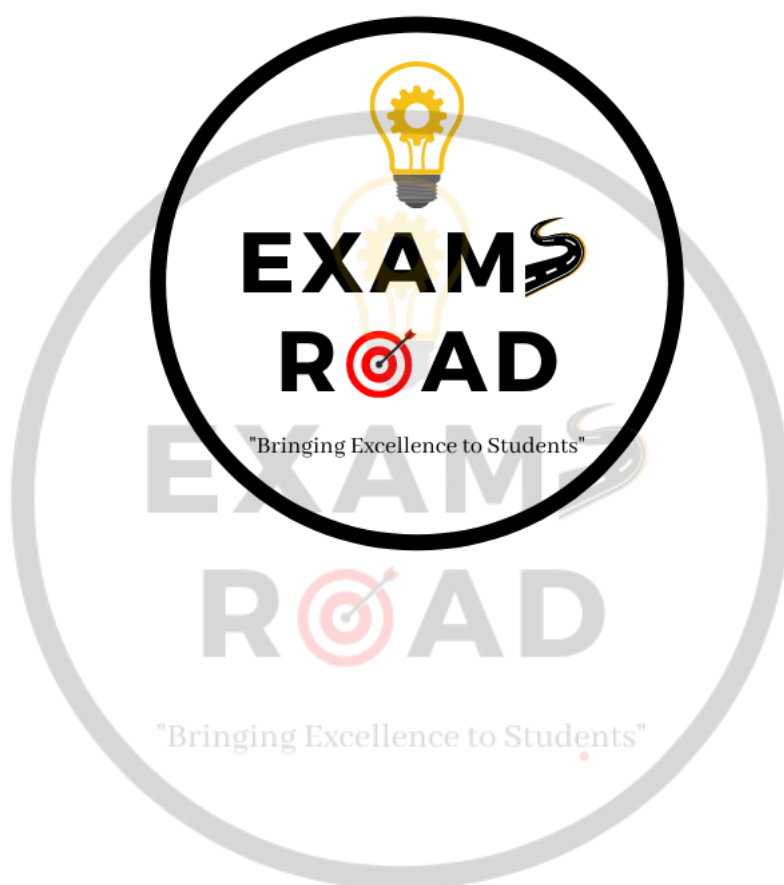


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Physical Chemistry



Basic Concepts of Chemistry

37

● Mole:

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

● Avogadro:

$$1 \text{ mole} \rightarrow 6.022 \times 10^{23}$$

● Volume:

$$1 \text{ mol (g)} \rightarrow 22.4 \text{ L}$$

● Vapour Density

$$\text{Molecular mass} = 2 \times \text{VD}$$

● Concentrations:

■ Molarity (moles of solute per litre solⁿ); $M = \frac{W}{m} \times \frac{1000}{V(\text{ml})}$ ↖ something just for ions.

■ Formality (moles of ion per litre solⁿ); $F = \frac{W}{m} \times \frac{1000}{V(\text{ml})}$

■ Molality (moles of solute per kg solvent); $m = \frac{W}{M} \times \frac{1000}{W'(\text{g})}$

■ Mole fraction x ; $x_i = \frac{n_i}{\sum n_i}$ $[\sum x_i = 1]$

■ Normality* (equivalents of solute per litre solⁿ); $N = \frac{W}{E} \times \frac{1000}{V(\text{ml})}$

■ $(\frac{W}{W})\%$ = $\frac{\text{mass of solute}}{\text{mass of sol}^n} \times 100$

■ $(\frac{W}{V})\%$ = $\frac{\text{mass of solute}}{\text{volume of sol}^n} \times 100$

■ $(\frac{V}{V})\%$ = $\frac{\text{volume of solute}}{\text{volume of sol}^n} \times 100$

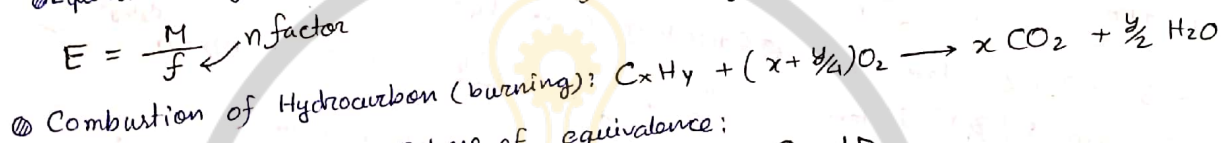
■ ppm (parts per million) = $\frac{\text{mass of solute}}{\text{mass of sol}^n} \times 10^6$

* Equivalent weight:

$$E = \frac{M}{f} \leftarrow n \text{ factor}$$

● n-factor:

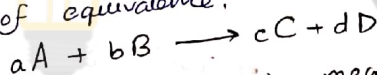
Acid	Base	Salt
n-f	Basicity	Acidity
		no of cations \times charge on cation



● Mili equivalents:

$$\text{meq} = NV = \frac{W}{E} \times 1000$$

● Law of equivalence:



Then, $\text{meq}(A) = \text{meq}(B) = \text{meq}(C) = \text{meq}(D)$ [When no limiting]

● Neutralisation:

■ Acid A + Acid B:

$$[H^+] = \frac{(NV)_A + (NV)_B}{V_A + V_B}$$

■ Base A + Base B:

$$[OH^-] = \frac{(NV)_A + (NV)_B}{V_A + V_B}$$

■ Acid A + Base B:

$$[H^+]/[OH^-] = \frac{|(NV)_A - (NV)_B|}{V_A + V_B}$$

■ If $(NV)_A = (NV)_B$,
[Neutral] pH = 7

● pH: $pH = -\log[H^+]$

$$pH + pOH = 14 \text{ [at } 25^\circ\text{C]}$$

■ Basic Medium:

▲ O_2 : Add water molecule^(same no) to the side where O_2 is excess.
Add double no of OH^- ion that of water to the other side.

▲ H_2 : Add H_2O molecule to the side where H is deficient.
Add equal no of OH^- on the other side.

● Balancing of O and H:

■ Acidic Medium:

▲ O_2 : Add H_2O molecule to the side where O_2 is deficient.

▲ H_2 : Add double no of H^+ ion where it is deficient.

● Balancing of Redox Reaction:

■ Identify oxdⁿ no of each element.

■ Identify oxdⁿ & reduction reaction.

■ Check if no of atoms of oxidant and reductant^{reduction} reaction.

■ Find out change in oxdⁿ no for oxdⁿ and redⁿ reaction.

■ Make the change in oxdⁿ no. equal for redⁿ rxn and oxdⁿ rxn by multiplying with a whole no.

Atomic Structure

● specific charge:

sp charge = $\frac{e}{m}$

● Energy of photon:

$E = h\nu = \frac{hc}{\lambda}$

● Bohr's Model:

■ Angular momentum:

$mvr = n \frac{h}{2\pi}$ [$n \rightarrow$ shell no.]

■ Transition Energy: $E_f - E_i$

$E_1 < E_2 < E_3 \dots < E_n$

$E_2 - E_1 > E_3 - E_2 > \dots > E_n - E_{n-1}$

■ Radius: $r = \frac{n^2 h^2}{4\pi^2 m k Z e^2} = 0.529 \frac{n^2}{Z} \text{ \AA}$

■ Velocity: $v = \frac{2\pi k e^2}{h} \left(\frac{Z}{n}\right) = 2.18 \times 10^8 \left(\frac{Z}{n}\right) \text{ cm s}^{-1}$

■ Total Energy:

$TE = KE + PE = \frac{1}{2} \frac{kZe^2}{r} - \frac{kZe^2}{r} = -\frac{1}{2} \frac{kZe^2}{r}$

$E_n = -13.6 \left(\frac{Z}{n}\right)^2 \text{ eV/atom}$
 $= -2.18 \times 10^{-18} \left(\frac{Z}{n}\right)^2 \text{ J/atom}$

■ Transition Energy:

$TE = 13.6 \times Z^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \text{ eV}$

■ Rydberg's Equation

$\frac{1}{\lambda} = \frac{2\pi^2 m k e^4}{h^3 c} \cdot Z^2 \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

$\frac{1}{\lambda} = R_H Z^2 \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

- Spectra:
- Lyman ($n_1 = 1$) [UV]
 - Balmer ($n_1 = 2$) [Visible]
 - Paschen ($n_1 = 3$) [IR]
 - Brackett ($n_1 = 4$) [IR]
 - Pfund ($n_1 = 5$) [IR]
 - Humphry ($n_1 = 6$) [IR]
 - Limit series ($n_1 = \infty$)

● Energy of an orbital:

$E_{n,l}$

■ If ($n+1$) is same, $E_{n,l}$

● Atomic Species:

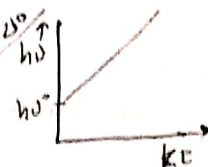
- Isotopes \rightarrow Same element, diff at. mass
- Isobars \rightarrow Same at. mass, diff element
- Isgotone \rightarrow Same no. of neutrons.
- Isodiaphors \rightarrow same isotopic no. [same $n-p$]
- Isotopic no = At. mass - 2 x at. no.
- Isoelectronic \rightarrow Same no. of e^-
- Isoster \rightarrow same no. of atoms and

● Photo-electric Effect: electrons are ejected from metal surface when being hit by photon with energy more than I.E

■ I.E/work function: $h\nu^0 / h\nu_{\lambda_0}$ [$\nu^0 \rightarrow$ Threshold frequency]

■ Einstein's Eqn on PE: $KE_e = h\nu - h\nu^0$

$E = \frac{1240 \text{ eV}}{\lambda}$ [λ is in nm (10^{-9} m)]



● De'broglie Hypothesis:

$\lambda = \frac{h}{m_0 v} = \frac{h}{\sqrt{2mKE}}$

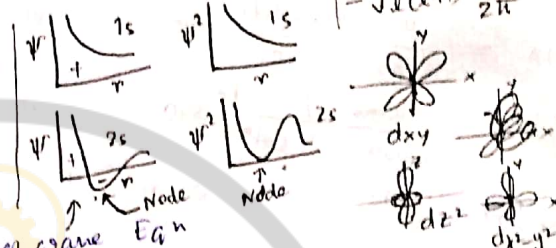
● Stopping Potential: $\lambda_s = \frac{h}{\sqrt{2mKE}}$ so that $KE_e = 0$

● Heisenberg's Uncertainty Principle

$\Delta x \times \Delta p \geq \frac{h}{4\pi}$

● Orbital angular momentum:

$= \sqrt{l(l+1)} \frac{h}{2\pi}$



● Schrodinger wave Eqn

$\frac{\partial^2 \psi}{\partial x^2} + \frac{\partial^2 \psi}{\partial y^2} + \frac{\partial^2 \psi}{\partial z^2} + \frac{8\pi^2 m}{h^2} (E - V)\psi = 0$

■ Solution gives value of n, l, m

■ ψ^2 = probability of find e^- [$1 > \psi^2 > 0$]

■ $\psi^2 = 0 \rightarrow$ node (no probability of electron)

■ Spherical nodes = $n - l - 1$

■ Angular nodes = l

■ Auf bau Principle: Particular order of filling orbitals

■ Pauli's Exclusion: n, l, m, s can't be same for 2 e^-

■ Hund's Rule: First all half filled the fill one by one.

■ Effective nuclear charge:

$Z_{eff} = Z - \sigma$



- Atomic Radius: (\rightarrow decreases) (\downarrow increases) Covalent Radius: $r_{cov} < r_{atom}$
- Metallic Radius: $r_{met} > r_{atom}$ Van-der-Waals radius: $r_{vdw} > \text{Everything}$ Iso-electronic: $Z \propto \frac{1}{size}$
- Shielding effect: $s > p > d > f$ Inert pair effect (Imp) Polarity: $\propto EN \text{ diff}$
- EN:

F	O	N	Cl	C	H
4	3.5	3	3	2.4	2.1

Chemical Bonding

- Covalent Bonding: VBT: Covalent bond formation takes place due to overlapping of half filled orbitals.
- Single bond: 1σ Double bond: $1\sigma, 1\pi$ Triple bond: $1\sigma, 2\pi$
- S orbitals always form σ bond. σ bond strength $> \pi$ bond strength due to higher extent of overlapping
- Types of overlapping: (s-s) / (s-p) / (p-p)

Hybridisation: process of intermixing of atomic orbitals having comparable energy to produce a new set of orbitals having similar properties.

- sp: Linear, 180°
- sp²: Trigonal Planar, 120°
- sp³: Tetrahedral, $109^\circ 28'$
- sp³d: TBP, $120^\circ/90^\circ$
- sp³d²: Octahedral, 90°
- sp³d³: Pentagonal bipyramidal

VSEPR Theory: $lp-lp > lp-bp > bp-bp$

Dipole Moment: (Due to EN diff):

$$\mu = e \times d \quad \mu_{net} = \sum \mu_i = \text{resultant}$$

Hybridisation	Geometry	LP	Shape	Ex
sp	linear	0	linear	CO ₂
sp ²	TP	0	TP	BF ₃
		1	Bent (\wedge)	SO ₂
		2	linear	NO ⁻
sp ³	Tetrahedral	0	Tetrahedral	CH ₄
		1	Pyramid	NH ₃
		2	Bent (\wedge)	H ₂ O
		3	linear	HOCl
sp ³ d	TBP	0	TBP	PCl ₅
		1	see-saw	SF ₄
		2	T-shape	BrCl ₃
		3	linear	XeF ₂
sp ³ d ²	Octahedral	0	Octahedral	SF ₆
		1	Square pyramid	XeOF ₄
		2	Square planar	XeF ₄
		3	T-shape	
sp ³ d ³	PBP	0	PBP	IF ₇
		1	Pentagonal py	XeF ₆

- DM of dichloro-benzene: ortho > meta > para
- DM of xylene: ortho > meta > para
- % ionic character = $\frac{\mu_{calc}}{\mu_{theo}} \times 100$ Unit of DM \rightarrow Debye (D)

Lattice Energy: $Na^+ + Cl^- \rightarrow NaCl + \Delta H$ Lattice energy.

Lattice Energy \propto charge density $\propto \frac{1}{size \text{ of ion}}$

- If solvation energy \geq Lattice energy \rightarrow then it dissolves.
- Polarity \propto solubility. Ionic bonds are non-directional.
- Covalent bonds are direction. MP, BP \rightarrow ionic > covalent.

Fajan's Rule: $\oplus + \ominus \rightarrow \oplus \ominus$ polarisation

- polarisation \propto covalent character.
- Polarising Power of cation $\propto \frac{1}{size}$
- Polarising power \rightarrow pseudoinert (Cu⁺) > inert core (K⁺)
- Polarisability \propto covalent character \propto size of anion.

Coordinate Covalent Bond: $H-N \rightarrow H$

Hydrogen Bond (Strong dipole-dipole): $H-F \cdots H-F$

- Intermolecular \rightarrow causes higher B.P. Doesn't occur in solid state
- That's why HF is liquid but HCl is gas.
- Without HB, BP \propto surface area \propto molar mass

Intramolecular HB: \rightarrow causes higher stability due to extra bond. Due to this, a ring formation (generally 5-6 no. ring) takes place.

London force: $\circ \circ$

Due to disturbance: $\circ \circ$

MOT [e \leq 14] $\sigma_{1s} < \sigma_{1s}^* < \sigma_{2s} < \sigma_{2s}^* < \pi_{2p_x} = \pi_{2p_y} < \sigma_{2p_z} < \pi_{2p_x}^* = \pi_{2p_y}^* < \sigma_{2p_z}^*$

[e \geq 14] $\sigma_{1s} < \sigma_{1s}^* < \sigma_{2s} < \sigma_{2s}^* < \sigma_{2p_z} < \pi_{2p_x} = \pi_{2p_y} < \pi_{2p_x}^* = \pi_{2p_y}^* < \sigma_{2p_z}^*$ nonmixing

Formed by linear combination of atomic orbitals \rightarrow constructive (same symmetry) \rightarrow Bonding Orbit
Destructive (oppo symmetry) \rightarrow Anti Bonding Orbit

Stability: BMO > AO > ABMO Hund's applicable

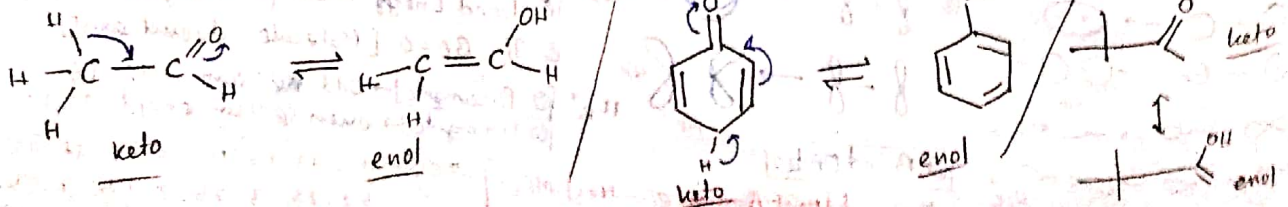
Bond Order: $B.O = \frac{1}{2} [No \text{ of Bonding } e^- - No \text{ of Antibonding } e^-]$

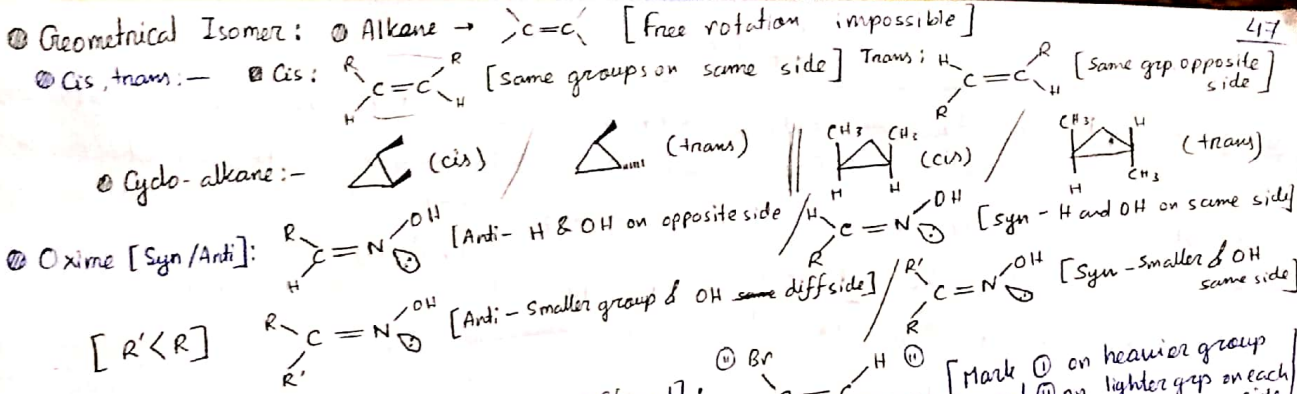
Gaseous State

- ① $T = t + 273$ ② $1L = 1dm^3 = 1000cc = 1000ml = 10^{-3}m^3$ ③ $1atm = 1bar = 760torr = 101325Pa$
- ④ **Boyle's Law:** $P \propto \frac{1}{V}$; $PV = \text{const.}$ ⑤ **Charles's Law:** $V \propto T$; $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ ⑥ **Avogadro's Law:** $V \propto n$; $\frac{V_1}{n_1} = \frac{V_2}{n_2}$
- ⑦ **Gay-Lussac's Law:** $P \propto T$; $P_T = \text{const.}$ ⑧ **Ideal Gas Equation:** $PV = nRT$; $PM = dRT$; $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$
- ⑨ **Dalton's Law of Partial Pressure:** $P_A = X_A P_{\text{Tot}}$ ⑩ $KE = \frac{3}{2} RT = \frac{3}{2} kT \text{ molecule}^{-1}$
- ⑪ **Graham's Law of Diffusion:** $\frac{r_1}{r_2} = \frac{P_1}{P_2} \sqrt{\frac{M_2}{M_1}}$
- ⑫ **Kinetic theory of Gas:** Pressure is due to collision of gas molecules with the walls of container.
- ⑬ **Postulates:**
 - ① Random motion causing collision with each other and wall.
 - ② collisions are perfectly elastic.
 - ③ Avg kinetic energy $\propto T$
 - ④ Volume of gas is negligible compared to volume of container.
 - ⑤ No intermolecular attraction or repulsion.
- ⑭ **KE:** $PV = \frac{1}{3} mn u^2$ ⑮ $U_{\text{avg}} = \sqrt{\frac{8RT}{\pi M}}$ ⑯ $U_{\text{rms}} = \sqrt{\frac{3RT}{M}}$ ⑰ $U_{\text{mp}} = \sqrt{\frac{2RT}{M}}$
- ⑱ **Ratio:** $U_{\text{rms}} : U_{\text{avg}} : U_{\text{mp}} = 1.22 : 1.12 : 1$
- ⑲ **Real Gas Equation (Vander-Wall):** $(P + \frac{an^2}{V^2})(V - nb) = nRT$
 - ① Excluded volume/co-volume: $b = 4N_A \times \frac{4}{3}\pi r^3$
 - ② Pressure correction = $\frac{an^2}{V^2} \propto (\frac{n}{V})^2$; $a \rightarrow$ signifies intermolecular attraction
- ⑳ **Compressibility factor (Z):**
 - ① at high Temp Press, $V_r > V_{\text{ideal}}$; $Z > 1 \leftarrow +ve$ deviation
 - ② at low Temp Press, $V_r < V_{\text{ideal}}$; $Z < 1 \leftarrow -ve$ deviation
- ㉑ **At very high Pressure:** $Z = 1 + \frac{Pb}{RT}$ ㉒ **At very Low Pressure:** $Z = 1 - \frac{a}{RTV}$
- ㉓ **Boyle's Temp:** $Z = 1$ for a long range of pressure.
- ㉔ **Real gas shows ideal behaviour at Low pressure and high temperature.**
- ㉕ **Inversion T = $\frac{2a}{Rb}$** ㉖ **Critical Temp = $\frac{8a}{27Rb}$** ㉗ **Critical Volume = $3b$** ㉘ **Critical Press = $\frac{8a}{27b^2}$**

Isomerism

- ① **Isomer:** Same molecular formula but different chemical properties.
- ② **Chain Isomers:** Different C-chain (\sim / \backslash)
- ③ **Functional Isomers:** Different functional group. [Alc \leftrightarrow Ether / Ald \leftrightarrow Ketone / CA \leftrightarrow Ester / Cyanide \leftrightarrow Isocyanide / Alkyne \leftrightarrow diene / Amine]
- ④ **General formulae:**
 - Alc \leftrightarrow Ether $\rightarrow C_n H_{2n+2} O$ ■ Ald, ketone $\rightarrow C_n H_{2n} O$
 - CA, Ester $\rightarrow C_n H_{2n} O_2$ ■ Cyanide - Isocyanide $\rightarrow C_n H_{2n-1} CN$ ■ Alkyne, diene $\rightarrow C_n H_{2n-2}$
 - Amine $\rightarrow C_n H_{2n+3} N$
- ⑤ **Alc, Ether \rightarrow [$\sim OH / \sim O \sim$]** ⑥ **Ald, ketone \rightarrow [$\sim CHO / \sim C=O$]** ⑦ **CA, Ester \rightarrow [$\sim COOH / \sim COO \sim$]**
- ⑧ **Cyan-Isocyan \rightarrow [$\sim CN / \sim NC$]** ⑨ **Alkyne-diene [$\sim \equiv / \sim =$]** ⑩ **Amine \rightarrow [$\sim N(1^\circ) / \sim N(2^\circ) / \sim N(3^\circ)$]**
- ⑪ **Positional Isomers:** Different position of func group. [occurs in Ether, Thioether, 2° amine, ket]
- ⑫ **Metamers:** Different no of C on both sides of func group. [occurs in Ether, Thioether, 2° amine, ket]
- ⑬ **Ring-chain isomers:** [DOU ≥ 1] [\sim / \square]
- ⑭ **Tautomers [Dynamic isomers as both exist in equilibrium]** [keto \rightleftharpoons enol]
- ⑮ **Conditions:**
 - There must be an sp^3 adjacent to the sp^2 Carbon
 - The sp^3 hybridised C must contain at least 1 H



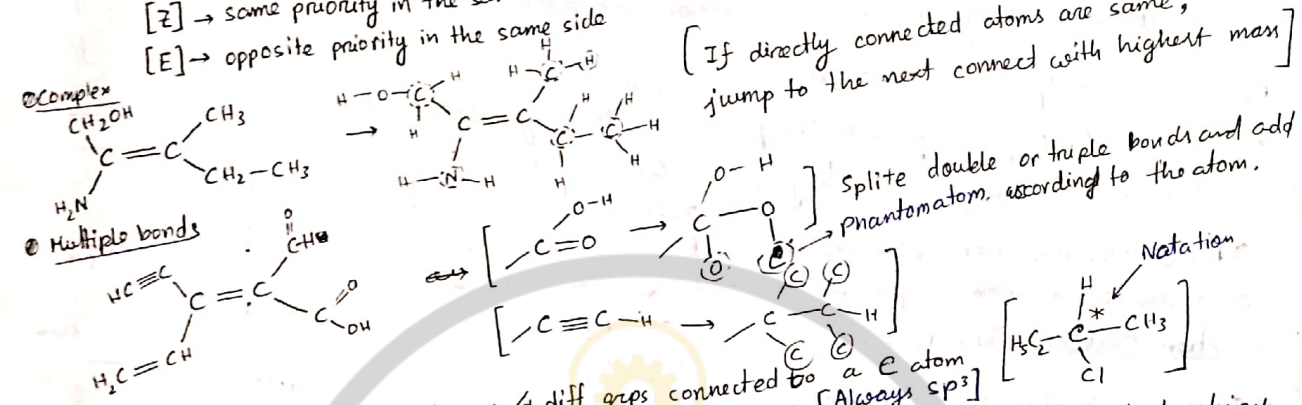


E-Z nomenclature [When All 4 grps are different]:

[Z] \rightarrow same priority in the same side

[E] \rightarrow opposite priority in the same side

Al. no \propto Priority Al. mass \propto Priority [In case of isotopes]



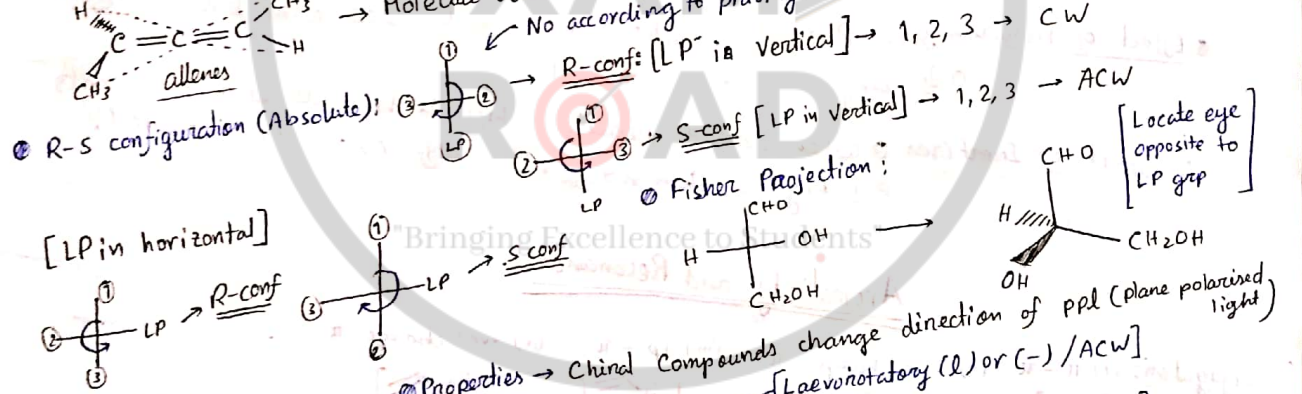
Optical Isomers: Chiral Carbon \rightarrow 4 diff grps connected to a C atom [Always sp^3]

CC have non-superimposable mirror image.

Molecule is chiral if it contains a chiral carbon

Molecule w/ more than 1 CC may or may not be chiral.

A molecule w/o plane of symmetry or centre of symmetry is always chiral.



2 chiral Carbons $\begin{array}{c} COOH \\ | \\ H \text{---} C \text{---} R \\ | \\ H \text{---} C \text{---} S \\ | \\ COOH \end{array} \rightarrow [2R, 3S]$

Properties \rightarrow Chiral Compounds change direction of PPL (plane polarised light)

ec/optically active \rightarrow [Laevorotatory (l) or (-) / ACW] \rightarrow [Dextrorotatory (d) or (+) / CW]

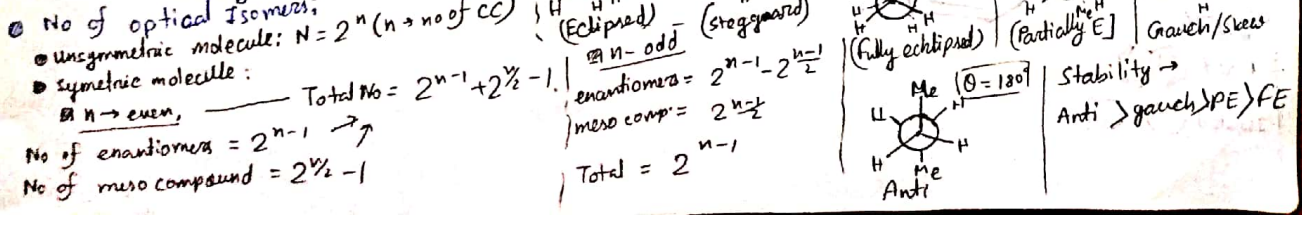
Enantiomers: stereo isomers which are non-superimposable mirror images of each other.

Diastereoisomers: stereo isomers which are not mirror images of each other.

Identical Isomer: stereo isomers \rightarrow superimposable mirror image.

Racemic Mixture: equimolar mixture of enantiomers (d & l) \rightarrow optically inactive.

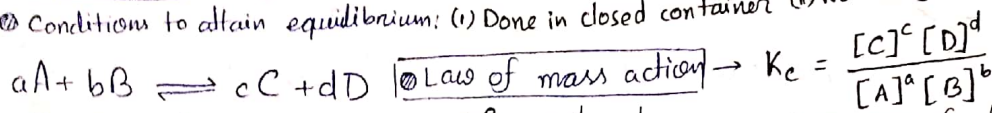
Meso Compound: Molecule w/ chiral Carbon as well as plane of symmetry i.e. not chiral. Phenomenon is called internal compensation. [They have identical Mirror Image]



Chemical Equilibrium

Equilibrium: Rate of fwd reaction = Rate of bkd rxn. Conditions: (i) $r_f = r_b$ (ii) $\Delta G = 0$ ($\Delta G = \Delta G_f - \Delta G_b$)

Conditions to attain equilibrium: (i) Done in closed container (ii) No loss of mass (iii) No precipitation occurs



In terms of pressure: $K_p = \frac{P_C^c \times P_D^d}{P_A^a \times P_B^b}$ P_A = Partial Pressure of A, $P_A = X_A P_T$

Relation: $K_p = K_c (RT)^{\Delta n_g}$

Reaction interted \rightarrow const. are reciprocal to previous. $[K = 1/K_0]$

Reaction Multiplied by n \rightarrow const. are raised to power n. $[K = K_0^n]$

Adding of Subtracting one or more reactions \rightarrow Adding \rightarrow multiply, subtract \rightarrow divide
 $[(i) + (ii) - (iii)] ; K = \frac{K_i \cdot K_{ii}}{K_{iii}}$

Pure solids and liquids doesn't involve in K terms.

Value of K changes only with temperature. Nothing else affects it.

Variation due to change in T: $\log \frac{K_2}{K_1} = \frac{\Delta H}{2.303 R} \left[\frac{1}{T_1} - \frac{1}{T_2} \right]$

For endothermic rxn, equilibrium const \uparrow with \uparrow in T

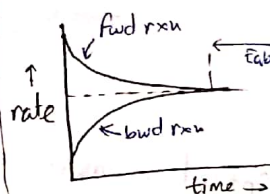
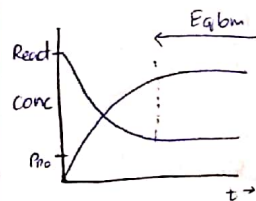
For exothermic rxn, equilibrium const. \downarrow with \uparrow in T.

Reaction Quotient: When $r_f \neq r_b$, $Q_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$

$Q_c = K_c$ | Reaction is at equilibrium

$Q_c > K_c$ | " shifts backwards

$Q_c < K_c$ | " shifts forward.



Le Chatelier's Principle: When eqbm is disturbed it shifts to the side where effect of disturbance is minimum.

Effect of Pressure: When P is \uparrow , reaction shifts to the side with lesser moles.

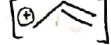
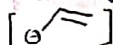
If $\Delta n_g = 0$, then no effect.

Effect of Inert Gas: When added, reaction shifts to the side with more moles.

If $\Delta n_g = 0$, then no effect. If $\Delta V = 0$, then no effect.

Aromaticity and Resonance

Conjugation: (i) $\pi - \pi$ (ii) $(-ve) \text{ charge} - \pi$ (iii) $lp - \pi$ (iv) $(+ve) \text{ charge} - \pi$



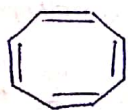
Condition \rightarrow All molecules should be sp^2 and possess sameily oriented p-orbital.

Aromatic Compound: cyclic, planer compounds containing odd pair $(4n+2)$ of delocalised π electrons.

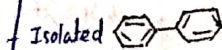
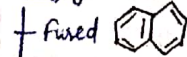
Antiaromatic: even pair $4n$ of π electrons. Non aromatic \rightarrow any other case.

Stability \rightarrow Aro $>$ Non aro $>$ Anti Aro. Compound w/ at least 1 benzene ring is aromatic

Aromatic Compound \rightarrow Benzenoids (Contains \odot ring)



Polycyclic



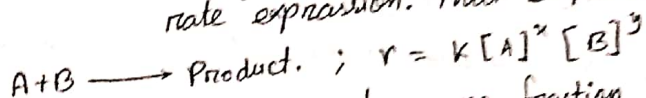
Non Benzenoids
(\odot ring absent)

Chemical Kinetics

● **Rate:** ● Average Rate = $\frac{-\Delta C[P]}{\Delta t} = \frac{\Delta C[P]}{\Delta t}$ ● Instantaneous Rate = $\frac{-dC[P]}{dt} = \frac{dC[P]}{dt}$ Rate of formation of P
 Representation: $aA + bB \rightarrow cC + dD \Rightarrow \text{Rate} = -\frac{1}{a} \frac{d[A]}{dt} = -\frac{1}{b} \frac{d[B]}{dt} = \frac{1}{c} \frac{d[C]}{dt} = \frac{1}{d} \frac{d[D]}{dt}$
 Rate of consumption of A

● **Rate Constant:** $A + B \rightarrow \text{Product}; r = k[A]^x[B]^y$
 $x \rightarrow$ order of reaction w.r.t. A
 $y \rightarrow$ order of reaction w.r.t. B
 $k \rightarrow$ rate const / specific rate

● **Order of Reaction:** When concentration of reactants are changed, how the rate of the reaction changes, is decided by the exponent of the concentration terms in the rate expression. That exponent terms are called order of the reaction.



● Order of a reaction can be zero, fraction, positive or negative.

● Order w.r.t. A = x , Order w.r.t. B = y , Overall order = $(x+y)$

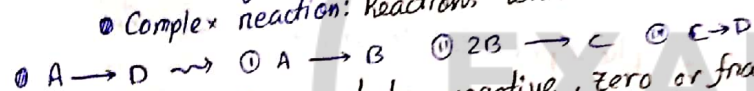
● 0 (zero) order means the concentration of the particular reactant does not affect rate.

● **Unit of Rate Constant:** $r = k[A]^n$; $k = r[A]^{-n} \rightarrow (\text{mol L}^{-1})\text{s}^{-1} (\text{mol L}^{-1})^{-n} = (\text{mol L}^{-1})^{1-n} \text{s}^{-1}$
 General unit: $(\text{mol L}^{-1})^{1-n} \text{s}^{-1}$ | $n \rightarrow$ overall order

● **Molecularity:** No of reactant molecules involved in an elementary reaction is called its molecularity.

● Elementary reaction: Reactions which can't be broken down further.

● Complex reaction: Reactions which can be broken down into elementary rxns.



● Molecularity cannot be negative, zero or fractional.

● Generally molecularity doesn't exceed 3.

● **Rate Determining Step:**

The slowest step is called RDS.

● **Integrated Rate Law:** $r = k[A]^n \Rightarrow -\frac{d[A]}{dt} = -k[A]^n \Rightarrow \left[\frac{1}{1-n} [A]^{1-n} \right]_{A_0}^{A_t} = -kt$
 (For $n \neq 1$)

$$\frac{1}{n-1} \left[[A]_t^{1-n} - [A_0]^{1-n} \right] = kt$$

● For 0th Order: $[A]_0 - [A]_t = kt$

● For 2nd Order: $\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$

For 1st Order: $kt = \ln \frac{[A]_0}{[A]_t} \Rightarrow t = \frac{2.303}{k} \log \frac{[A]_0}{[A]_t} \Rightarrow [A]_t = [A]_0 e^{-kt}$

● **Half-life Period:** $[A]_t = \frac{1}{2} [A]_0 \rightarrow$ time taken.

● 0th Order: $t_{1/2} = \frac{[A]_0}{2k}$

1st Order:

$$t_{1/2} = \frac{0.693}{k}$$

2nd order: $t_{1/2} = \frac{2}{[A]_0 k}$

● nth order:

$$t_{1/2} = \frac{[A]_0^{1-n}}{k(1-n)} [2^{1-n} - 1]$$

no of half lives

Thermodynamics

● **System** Part of universe where something is done.

- Open System: Exchange of mass and heat.
- Closed System: Exchange of energy not mass
- Isolated System: Exchange of none (mass or energy)

● **State function**: Only depends upon initial and final state [e.g. $S, q, \Delta G$]

● **Path function**: Depends on path followed [$\Delta Q, W$]

● **Extensive property**: Dependent on amount or mass

● **Process** Method of doing something.

- **Reversible**: Change in energy step is negligible, system and surroundings stay in eqbm [Time \rightarrow infinite]
- **Irreversible**: Changes are not negligible, system and surroundings not in eqbm.

● **Intensive property**: Independent on amount or mass

● **1st Law of Thermodynamics** Energy can neither be created nor destroyed it can be converted across form

$$\Delta U = Q + W \quad [\Delta U \rightarrow \text{change in internal energy}, Q \rightarrow \text{Heat given to the system}, W \rightarrow \text{work done on the system}]$$

● Work in terms of P-V: $W = -P_{\text{ext}} \Delta V$ ● Work in case of a reversible & Isothermal process:

● Work in isochoric process: ($\Delta V = 0$)

$$\therefore W = 0$$

● Work in free expansion: ($P_{\text{ext}} = 0$)

$$\therefore W = 0$$

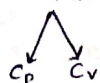
● Work in case of irreversible process:

$$W = -P_{\text{ext}}(V_2 - V_1)$$

● **Heat Capacity** Amount of heat required to raise the temp by 1°C , for a given amount of mass

● **Specific Heat capacity**: Heat required to raise the temp by 1°C , for unit mass.

● **Molar Heat capacity**: Heat required to raise the temp by 1°C , for one mole



$$Q = ms\Delta T$$

● **Cp** Molar heat capacity at const. pressure. [Heat provided is utilised to raise the temp. as well as to do work, Thus more heat is needed]

$$C_p = \left(\frac{\partial H}{\partial T}\right)_P = \left(\frac{\Delta H}{\Delta T}\right)_P$$

● **Cv** Molar heat capacity at const. volume [Heat provided is utilised only to raise temp, Thus less heat is needed]



$$C_v = \left(\frac{\partial U}{\partial T}\right)_V = \left(\frac{\Delta U}{\Delta T}\right)_V$$

$$C_p > C_v$$

● **Enthalpy** $\Delta H = \Delta U + P\Delta V + V\Delta P = \Delta U + nRT$

● At const T: $\Delta H - \Delta U = \Delta n_g RT$

● No of moles when $\Delta T \neq 0$

Enthalpy change at const. press.

Enthalpy change at const. volume.

● **Work for adiabatic reversible process**

$$W = \frac{R}{\gamma - 1} (T_2 - T_1)$$

$$\Delta H = \Delta U + nR\Delta T$$

$$\frac{\Delta H}{\Delta T} = \frac{\Delta U}{\Delta T} + nR$$

$$C_p - C_v = nR$$

● **Poisson Ratio**

For 1 mole, \rightarrow

$$C_v = \frac{R}{\gamma - 1}$$

● **Entropy** measure of disorderness [State function, Extensive property]

$$\Delta S = C_v \ln \left(\frac{T_2}{T_1}\right) + R \ln \left(\frac{V_2}{V_1}\right) \rightarrow \text{for } \Delta T \neq 0: \Delta S = nR \ln \frac{V_2}{V_1} = nR \ln \frac{P_1}{P_2}$$

$$\rightarrow \text{for } \Delta V \neq 0: \Delta S = C_p \ln \left(\frac{T_2}{T_1}\right)$$

$$\text{Relations for adiabatic process: } \bullet PV^\gamma = \text{const.} \bullet TV^{\gamma-1} = \text{const.} \bullet TP^{(1-\gamma)} = \text{const.}$$

● **2nd Law of Thermodynamics** for an isolated system, change in entropy is always positive.

$$\Delta G = \Delta H - T\Delta S$$

- For spontaneous process $\rightarrow \Delta S_{\text{tot}} > 0$; $\Delta G_{\text{sys}} < 0$
- For equilibrium $\rightarrow \Delta S_{\text{tot}} = 0$; $\Delta G = 0$

● **Free energy** Amount of energy stored in a system that can be converted into useful work.

$$\Delta G = \Delta G^\circ + RT \ln Q$$

\rightarrow for Equilibrium:

$$\Delta G^\circ = -RT \ln K$$

$$K = e^{-\frac{\Delta G^\circ}{RT}}$$

● **ΔG° for reaction**

$$\Delta G^\circ_{\text{rxn}} = \sum \Delta G^\circ_{\text{product}} - \sum \Delta G^\circ_{\text{reactant}}$$

$$\rightarrow aA + bB \rightarrow cC + dD: \Delta G^\circ_{\text{rxn}} = c\Delta G^\circ(C) + d\Delta G^\circ(D) - a\Delta G^\circ(A) - b\Delta G^\circ(B)$$

$$\Delta S^\circ_{\text{rxn}} = c\Delta S^\circ(C) + d\Delta S^\circ(D) - a\Delta S^\circ(A) - b\Delta S^\circ(B)$$

Electrolytes Which

Strong Electrolyte

Weak Electrolyte

pH Scale # pH =

base # pH =

neutral # pH =

acid # Kw =

pH can

Calculation of pH

$$t \alpha = \sqrt{\frac{K_a}{C}} \quad \# [OH^-]$$

$$pOH = \frac{1}{2} [pK_a - \log C]$$

pH calculation for

alt $\rightarrow SA + S$

$\rightarrow SB + V$

$\rightarrow WB + S$

$\rightarrow WA + W$

$$K_h = \frac{K_w}{K_b}$$

$$[H^+] = C\alpha = C\sqrt{\frac{K_a}{C}}$$

$$pH = \frac{1}{2} [pK_a - \log C]$$

$$= 7 - \frac{1}{2}$$

$$pH = 7 + \frac{1}{2}$$

$$pH = 7 + \frac{1}{2}$$

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Ionic Equilibrium

is negligible,
m [Time → infinite]
ible, system and

amount on mass.

noted across forms.

→ work done on the sys.

rocess:

1/2)

1/2)

ount of mass.

1.

b

to raise the

to raise

= ΔU + nRT

= $\frac{\Delta U}{\Delta T} + nR$

- C_v = nR

dq_{rev}

= nR ln $\frac{P_1}{P_2}$

1)

positive.

eful work.

adant

0°(B)

0°(B)

Electrolytes Which can dissociate into ions in molten or aqueous state.

Strong Electrolytes (Dissociates 100%) e.g. HCl, H₂SO₄, HNO₃

Weak Electrolytes: (Dissociates 100%) e.g. HCN, H₂SO₄, HNO₃

pH Scale # pH = -log₁₀ [H⁺], # pOH = -log [OH⁻]

pH + pOH = 14 [At 25°C] Ionic product of H₂O

pH + pOH = pK_w # K_w = [H⁺][OH⁻]

K_w = 10⁻¹⁴ [At 25°C] # K_w ↑ with T ↑

pH can be negative

Calculation of pH for weak base

α = $\sqrt{\frac{K_b}{C}}$ # [OH⁻] = Cα = $\sqrt{CK_b}$

pOH = $\frac{1}{2} [pK_b - \log C]$

pH calculation for Salts

Salt → SA + SB (No Hydrolysis) [pH = 7]

→ SB + WA (Hydrolysis) / Cationic Hydrolysis [pH > 7]

→ WB + SA (Hydrolysis) / anionic Hydrolysis [pH < 7]

→ WA + WB (Hydrolysis) [pH = 7]

K_h = $\frac{K_w}{K_b}$ # for hydrolysis (α < 10%)

→ α = $\sqrt{\frac{K_h}{C}}$

[H⁺] = Cα = $C\sqrt{\frac{K_h}{C}}$ = $\sqrt{CK_h}$ = $\sqrt{C \cdot \frac{K_w}{K_b}}$

SA + WB # pH = $\frac{1}{2} [pK_w - pK_b - \log C]$

= $7 - \frac{1}{2} [pK_b + \log C]$ salt concentration

SB + WA # pH = $7 + \frac{1}{2} (pK_a + \log C)$

WA + WB # pH = $7 + \frac{1}{2} [pK_a - pK_b]$

All these formulas are used when millieq of both acid and base is same.

If we add H⁺ to buffer: CH₃COO⁻ + H⁺ → CH₃COOH

The above reaction very fast (K = $\frac{1}{1.85 \times 10^{-5}}$). The added H⁺ immediately converted into weak acid and blocks the acids dissociation. ∴ pH ≈ 0

If we add OH⁻ to buffer: CH₃COOH + OH⁻ → CH₃COO⁻ + H₂O

CH₃COOH + H₂O ⇌ CH₃COOH + OH⁻ ← This hydrolysis rxn is really slow, amount of OH⁻ furnished is negligible.

pH of Buffer solution doesn't change on dilution.

pH of salt Buffer is independent of salts conc.

pH calculation for a mixture of two weak acid

HA ⇌ H⁺ + A⁻ | HA' ⇌ H⁺ + A'⁻

C(1-α₁) α₁ α₁ | C(1-α₂) α₂ α₂

K_{a1} = $\frac{\alpha_1(\alpha_1 + \alpha_2)}{(1 - \alpha_1)}$ K_{a2} = $\frac{\alpha_2(\alpha_1 + \alpha_2)}{(1 - \alpha_2)}$

To be Continued

Acid base

Arrhenius Acid → H⁺ in aq. Base → OH⁻ in aq.

Lewis e⁻ pair donor (Acid) e⁻ pair acceptor (base)

Bronsted proton donor (Base) proton acceptor (Acid)

HI > HBr > HCl [Acid strength]

pH for mixture of strong acid & base

pSA + SA → [H⁺] = $\frac{N_1V_1 + N_2V_2}{V_1 + V_2}$

pSB + SB → [OH⁻] = $\frac{N_1V_1 + N_2V_2}{V_1 + V_2}$

SA + SB → [H⁺] = $\frac{N_1V_1 - N_2V_2}{V_1 + V_2}$

Buffer Solution

A solution which can resist the change in pH when a small amount of acid or base added to it is called a buffer solution.

Acid Buffer: (WA + salt of WA & SB)

e.g. CH₃COOH + CH₃COONa / HCN + NaCN

Basic Buffer: (WB + salt of WB & SA)

e.g. NH₄OH + NH₄Cl

Neutral Buffer / Salt Buffer: (salt of WA & WB)

Acid Buffer → pH = pK_a + log $\frac{[salt]}{[acid]}$

Basic Buffer → pOH = pK_b + log $\frac{[salt]}{[base]}$

Salt Buffer → pH = 7 + $\frac{1}{2} [pK_a - pK_b]$

Common Ion effect

In presence of a common ion, the degree of dissociation of an weak electrolyte decreases (WA/WB)

Buffer action → CH₃COOH + CH₃COONa (strong electrolyte)

(weak electrolyte)

CH₃COO⁻ + H⁺ → CH₃COOH

CH₃COO⁻ + Na⁺ → CH₃COONa

due to common ion effect, this forms very little.

Solubility & Solubility Product

g/L → Amount of solute is grams dissolved to prepare 1L saturated solution.

Molar Solubility → No of moles of solute dissolved to prepare 1L saturated solution.

A_xB_y ⇌ xA^{y+} + yB^{x-}

K_{sp} = [A^{y+}]^x [B^{x-}]^y

= x^x y^y s^{x+y}

Q_{sp} > K_{sp} → Then the equation will move backwards, precipitation occurs (supersaturated)

Q_{sp} < K_{sp} → The equation will move forward (unsaturated)

Q_{sp} = K_{sp} (saturated)

In presence of common ion, the solubility decreases.

antibiotics

Boron family (Group 13)

B	[He] 2s ² 2p ¹
Al	[Ne] 3s ² 2p ¹
Ga	[Ar] 3d ¹⁰ 4s ² 4p ¹
In	[Kr] 4d ¹⁰ 5s ² 5p ¹
Tl	[Xe] 4f ¹⁴ 5d ¹⁰ 6s ² 6p ¹

Valance shell electronic configuration $\rightarrow ns^2 np^1$

Element	Ox no.
B	+3
Al	+3
Ga	+1, +3
In	
Tl	

Stability: $Ga^+ < In^+ < Tl^+$
 $Ga^{3+} > In^{3+} > Tl^{3+}$

Due to inert pair effect.
 Down the group, lower ox no. is more stable for heavier p block elements (containing d or f orbitals)

Stability: $Ga^{3+} > Ga^+$
 $Ga^+ \rightarrow Ga^{3+} + 2e^-$

Stability $Tl^+ > Tl^{3+}$
 $(Tl^{3+}) + 2e^- \rightarrow Tl^+$
 oxidising agent.

Atomic Radius
 $B < Ga < Al < In < Tl$
 Exception due to d orbital.

IE $B > Ga > Al > In > Tl$

Melting point
 B Al Ga In Tl
 decrease increase
 Ga has lowest m.p among the group 13 elements due to its unusual crystalline structure.

All Group 13 elements form halides reacting with halogen
 $B + X_2 \rightarrow BX_3$ - covalent
 $Al + X_2 \rightarrow AlX_3$
 $Ga + X_2 \rightarrow GaX_3$
 $In + X_2 \rightarrow InX_3$
 $Tl + X_2 \rightarrow TlX_3$ - ionic

Exception

$N(CH_3) \rightarrow$ Pyramidal (sp^3) [No d' orbital; No back bonding]

$N(SiH_3) \rightarrow$ Trigonal (sp^2) [d' orbital present]

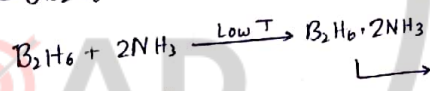
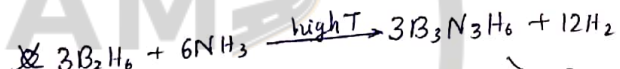
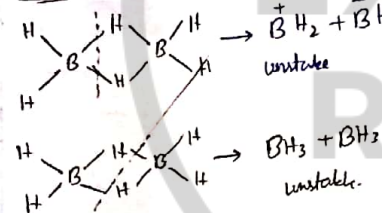
Borane $[BH_3]$

Diborane $[B_2H_6]$



Bonding Hydrogen and Terminal H lie in 1 planes

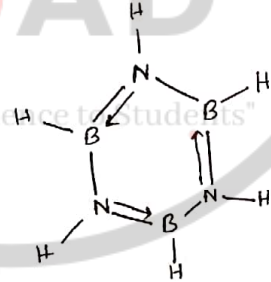
Reactions



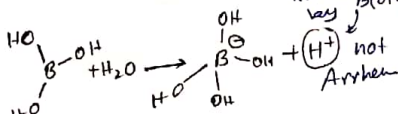
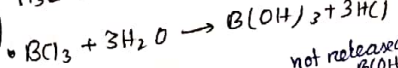
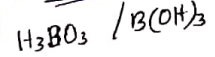
Borazine / Inorganic benzene.

[Stability: Benzene > Borazine]

Bond breaking will take place in presence of nucleophile. Due to steric hindrance, larger nucleophiles (\wedge in) can break to form BH_2 and BH_4

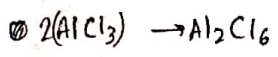
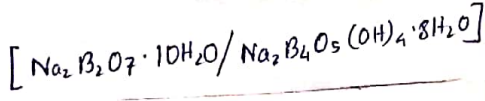
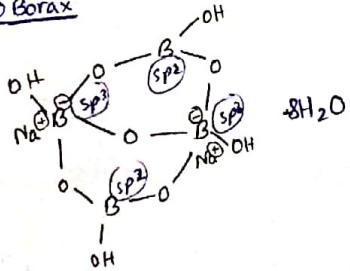


Boric Acid

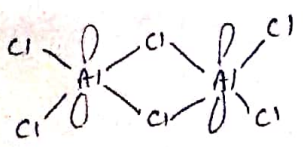


not giving H^+ but accepting OH makes it an acid.

Borax



$2AlF_3 \rightarrow Al_2F_6$
 [Doesn't form]
 cos F is too small

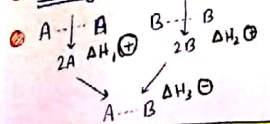


$B(OH)_3$	$Al(OH)_3, Ga(OH)_3$	$In(OH)_3, Tl(OH)_3$
acidic	amphoteric	basic

Liquid Solution

95

Binary solution: Solution with only 2 constituent



$$\Delta H_{sol} = (\Delta H_1 + \Delta H_2) + \Delta H_3$$

- $\Delta H_{sol} = 0 \rightarrow$ Ideal solⁿ
- $\Delta H_{sol} < 0 \rightarrow$ -ve deviation
- $\Delta H_{sol} > 0 \rightarrow$ +ve deviation

non volatile solute + volatile solvent [Salt/sugar in water]
 volatile solute + volatile solvent [Alcohol in water]

When V_1 ml of A and V_2 ml of B are mixed, the volume of the solution can be $<$, $=$ or $>$ ($V_1 + V_2$)

- $\Delta V < 0$ (-ve deviation) [Due to strong bond formation]
- $\Delta V = 0$ [Ideal]
- $\Delta V > 0$ (+ve deviation) [Due to strong break bond formation]

Ideal Solution:
 $\rightarrow \Delta H = 0$ (change of solⁿ)
 $\rightarrow \Delta V = 0$ (change in volume)
 \rightarrow Obeys Raoult's Law

- If $\Delta H < 0 \rightarrow$ also $\Delta V < 0$
- If $\Delta H > 0 \rightarrow$ also $\Delta V > 0$
- If $\Delta H = 0 \rightarrow$ also $\Delta V = 0$

Raoult's Law: Vapour pressure of the solution is directly proportional to the mole fraction of the volatile solute.

$$P_{sol} \propto X_{\text{volatile compound}}$$

Vapour Pressure: Pressure exerted by the vapour at equilibrium is called V.P.

Colligative Property: Only depends on amount or collection not quality or nature.

$$V.P. \propto \frac{1}{b.p.}$$

$$X_{\text{solute}} = \frac{P^0 - P^s}{P^0}$$

Relative lowering of V.P.

$$\text{Lowering of V.P.} = P^0 - P^s$$

For a dilute solution ($n_A + n_B \approx n_B$)

$$\frac{P^0 - P^s}{P^0} = \frac{W_A}{M_A} \times \frac{M_B}{W_B}$$

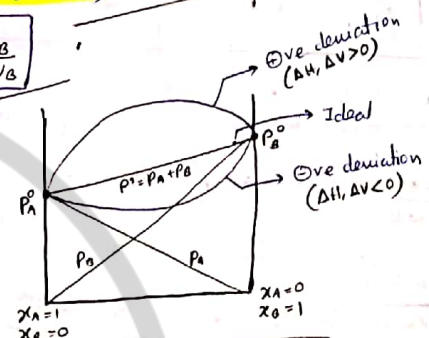
When Both are volatile: $A \rightarrow$ solute, $B \rightarrow$ solvent

$$P_A = P_A^0 X_A \rightarrow P_A^s = P_A + P_B$$

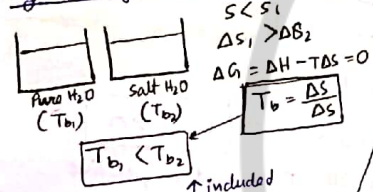
$$P_B = P_B^0 X_B \rightarrow P^s = P_A^0 X_A + P_B^0 X_B$$

$$P^s = P_A^0 (1 - X_B) + P_B^0 X_B$$

$$P^s = P_A^0 + (P_B^0 - P_A^0) X_B$$



By Thermodynamics:



$$S < S_1$$

$$\Delta S_1 > \Delta S_2$$

$$\Delta G = \Delta H - T\Delta S = 0$$

$$T_b = \frac{\Delta S}{\Delta S}$$

Vant Hoff factor (i):

$$i = \frac{\text{No of moles of ions in the solution}}{\text{no of moles of solute dissolved}}$$

- $i = 1 \rightarrow$ No dissociation, no association
- $i > 1 \rightarrow$ dissociation
- $i < 1 \rightarrow$ Association

Molality:

$$m = \frac{n_A}{n_B \times M_B} \times 1000$$

$$= \frac{X_A}{X_B \times M_B} \times 1000$$

Elevation in B.P.

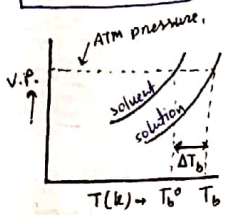
$$\Delta T_b = T_b^s - T_b^0 = \text{+ve}$$

$$\Delta T_b = i \cdot K_b \cdot m$$

i = vant Hoff factor
 K_b = ebullioscopic const. for the solvent
 m \rightarrow molality

$$K_b = \frac{R T_b^2 M}{1000 \Delta H_{\text{vap}}}$$

\rightarrow B.P. of solvent in K
 \rightarrow Molar mass of solvent
 \rightarrow Enthalpy of vap of solvent/mole



$$K_b = \frac{R T_b^2}{\Delta H_{\text{vap}} \times 1000}$$

latent heat of vap

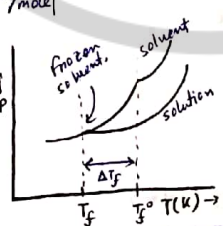
Depression in F.P.

$$\Delta T_f = T_f^0 - T_f^s = \text{+ve}$$

$$\Delta T_f = i K_f m$$

$$K_f = \frac{R T_f^2 M}{1000 \Delta H_{\text{fusion}}}$$

[Symbols have usual meanings]



$$K_f = \frac{R T_f^2}{\Delta H_{\text{fusion}} \times 1000}$$

latent heat of fusion

Osmotic Pressure:

\rightarrow Osmosis: When 2 solutions of different concentrations are separated with a semi membrane, solvent particles move from its higher concentration site to its lower concentration site [from lower conc solution to higher]. This phenomenon is called Osmosis.

\rightarrow Osmotic pressure: The minimum pressure needed to be applied on the higher concentration site of the solution to stop osmosis. If pressure is more than osmotic pressure, reverse osmosis takes place (RO). Used in purifying water.

$$\text{Osmotic pressure, } \pi = \frac{W_B R T}{M_B V}$$

\rightarrow Temp
 \rightarrow core
 \rightarrow V.T. fact
 \rightarrow molecule weight of solute

Osmotic pressure of mixture:

\rightarrow Two solutions of same substance:

$$\pi_1 V_1 + \pi_2 V_2 = \pi_R (V_1 + V_2)$$

\rightarrow n_1 and n_2 moles of two diff solutes present in V_1 and V_2 volumes.

$$\pi_R = \pi_1 + \pi_2 = \frac{n_1 i_1 R T}{(V_1 + V_2)} + \frac{n_2 i_2 R T}{(V_1 + V_2)}$$

$$\pi_R = \frac{(n_1 i_1 + n_2 i_2) R T}{(V_1 + V_2)}$$

Types of Solution:

\rightarrow **Isotonic:** Solution with same osmotic pressure on both sides

\rightarrow **Hyper tonic:** Solutions of equimolar concentration at the same Temp, if a solution has more O.P. than the other, it's Hyper tonic

Hypo tonic: The solution which has lesser O.P. than the other one is called Hypo tonic.

★ When colligative properties are used in case of dilute solution, we can consider, molality \approx molarity.

Relationship between different colligative properties:

O.P & Relative lowering of V.P: $\pi = \frac{dRT}{Mw_1} \left[\frac{p^0 - p_s}{p^0} \right]$

$d(\text{density}) = \frac{W_1}{V}$

EBP & RVP: $\Delta T_b = \frac{1000 \times K_b}{Mw_1} \left(\frac{p^0 - p_s}{p^0} \right)$

DEP & RVP: $\Delta T_f = \frac{1000 \times K_f}{Mw_1} \left(\frac{p^0 - p_s}{p^0} \right)$

O.P & Elevation in B.P:

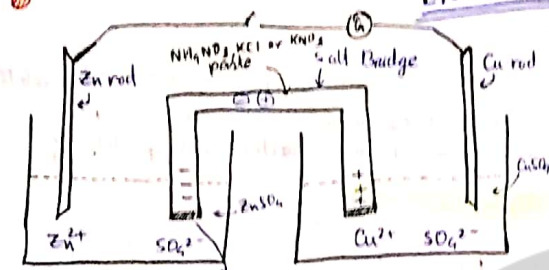
$\pi = \frac{\Delta T_b R T \times d}{K_b \times 1000}$

O.P & Depression in F.P:

$\pi = \frac{\Delta T_f \times d R T}{K_f \times 1000}$

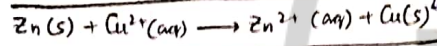
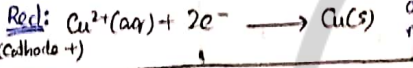
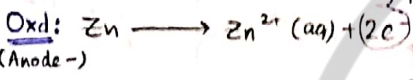
$d(\text{density}) = \frac{W_1}{V}$

ELECTROCHEMISTRY



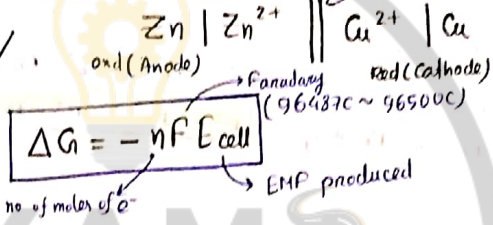
→ When switch is closed, current starts flowing.
 → In the salt bridge, salt cations and anions have comparable ionic mobility i.e. size
 → Zn rod becomes thinner / Cu rod becomes thicker
 Zn is more electropositive than Cu, Zn has higher tendency of losing electron i.e. oxidation potential than Cu
 Cu has higher reduction potential

Doesn't allow the paste to pass but ions



overall cell reaction

Short-hand Notation:



Standard (for molarity = 1)

$\Delta G^\circ = -nF E^\circ_{cell}$

→ $\Delta G = \Delta G^\circ + RT \ln Q$

Nernst Equation

$E_{cell} = E^\circ_{cell} - \frac{RT}{nF} \ln \frac{[Zn^{2+}]}{[Cu^{2+}]}$

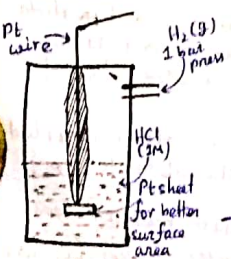
$E^\circ_{cell} = E^\circ_{red}(\text{cathode}) - E^\circ_{red}(\text{anode})$

for spontaneous reaction:

$\Delta G < 0 \rightarrow -nF E_{cell} < 0 \rightarrow E_{cell} > 0$

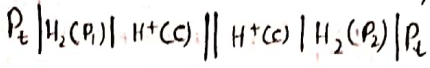
for Galvanic cell / Daniel Cell, E_{cell} must be positive

SHE (Standard Hydrogen Electrode):



If redⁿ potential of other electrode is more, H_2 will go through.

→ Pt wire works as electron carrier.



$Q = \left(\frac{P_{H_2}(P_2)}{P_{H_2}(P_1)} \right)^{1/2}$

Std reduction potential of $H = 0$ (by convention)

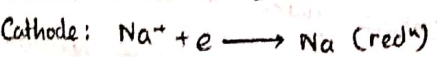
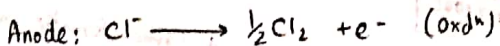
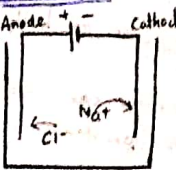
$E_{cell} = E^\circ_{cell} - \frac{RT}{nF} \ln Q$

→ $E_{cell} = - \frac{RT}{nF} \ln Q$
 as $E^\circ_{cell} = 0$

As $P_1 > P_2$
 $Q_1 = -ve$
 $\ln Q = -ve$
 when $P_1 = P_2$ / $E_{cell} = 0$

→ Where pressure is high, H_2 needs to be reduced, oxidation occurs
 → low, oxidise, reduction occurs

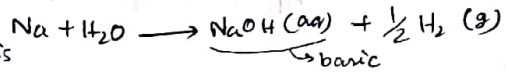
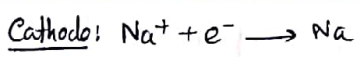
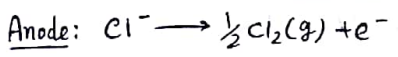
Electrolysis



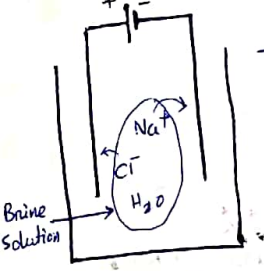
→ In galvanic cell, anode was negatively charged
 → In electrolytic cell, anode is positively charged

Def: The process of separation of ions by passing electric current through its aqueous solution or molten solution in an electrolytic cell is called electrolysis.

If they are aqueous



→ Brine solution is neutral by after electrolysis the solution becomes basic



1st Law: $W \propto Q \rightarrow W = ZQ$ | $Q = it$ | Electrochemical equivalent
 Amount of charge | time | current
 ↓ | ↓ | ↓
 charge | current |
 $\therefore W = Zit$

When 1C charge is passed thru the solⁿ amount deposited is called E_{CE} [$C=1, W=Z$]

Faraday's law of electrolysis:

Amount deposited (g) at electrode is directly proportional to the charge (C) passing thru its solution

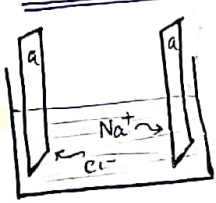
No. of g eq^t $\left(\frac{W}{E} \right) = \frac{Q}{F} \rightarrow$ comparing 1st and Law, $Z = \frac{E}{F}$

2nd Law:

↓ 2 or more E cell connected in series to a single, then ratio of the amount deposited is equal to the ratio of equivalent masses.

$\frac{W_A}{E_A} = \frac{Q}{F}, \frac{W_B}{E_B} = \frac{Q}{F}, \frac{W_A}{W_B} = \frac{E_A}{E_B}$

Conductance and Conductivity:



Conductance
 $R \propto l \rightarrow R \propto \frac{l}{a}$
 $R = \rho \frac{l}{a}$
 $\frac{1}{R} = \frac{1}{\rho} \frac{a}{l}$
 resistivity
 $G = K \frac{a}{l}$
 conductivity

$G = K \cdot \frac{A}{l}$
 $G l^2 = K V$
 If $l = 1\text{m}$ or 1cm
 $G = K V$

→ If a solution contains 1 mole of electrolyte, its conductance is molar conductance ($\Delta_m = K \cdot V$)
 → If a solution contains 1 g-eq of electrolyte, ---
 --- equivalent conductance. (Δ_{eq})

$\Delta_{\text{eq}} = \Delta_m \times n\text{-factor}$

