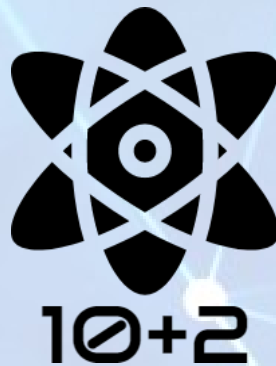


10+2 PCM NOTES

BY

JOYOSHISH SAHA

(PDF version handwritten notes of Maths, Physics and Chemistry for 10+2 competitive exams like JEE Main, WBJEE, NEST, IISER Entrance Exam, CUCET, AIPMT, JIPMER, EAMCET etc.)



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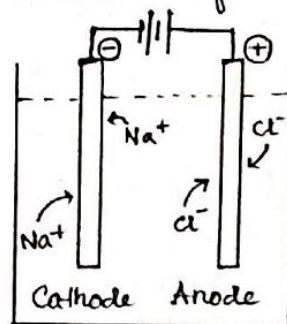
With best wishes from Joyoshish Saha

- * Study of production of electricity from energy released during spontaneous chemical reactions & use of electrical energy to bring about non-spontaneous chemical transformations.
- * If electrical energy is used to bring about a chemical change in a cell, then the cell is called an electrolytic cell.

In a cell, if chemical changes produce electricity, then the cell is called electrochemical cell or Voltaic cell or Galvanic cell.

- * **Electrolytic Cell:** An electrolytic cell is an arrangement in which electricity is conducted through a solution or a molten state by the movement of ions.

- The ability of cation to move towards the negative cathode & get reduced depends upon the size, mass, positive charge etc.



- A cation with a higher value of Standard Reduction Potential (SRP) would get reduced in preference to a cation with a lower value of SRP. (F_2 has highest SRP (2.87V) & Li has lowest).

- Ability of a cation to get reduced or an anion to be oxidised not only depends on their SRP's, but also depends on their concentrations.

* Faraday's Laws of Electrolysis:

- First law: $w \propto q$ [$w \rightarrow$ mass of substance deposited or liberated at any electrode,
 $\Rightarrow w = Zq = Zit$
 $Z \rightarrow$ electrochemical equivalent. $q \rightarrow$ amount of charge passed]. $Z = \frac{\text{eq. wt of substance}}{96500}$

- Second Law: $w \propto E$ [$w \rightarrow$ mass of substance in gram, $E \rightarrow$ equivalent weight in gram].
 $\Rightarrow \frac{w_1}{E_1} = \frac{w_2}{E_2}$

• 1 Faraday: The charge possessed by 1 mole of electrons. (96487 C).

1 Faraday of charge liberates 1 gm equivalent weight of substance.

* Electrolytic Conductance: $\sigma = \frac{1}{R}$ [ohm^{-1}].

Magnitude of the conductance depends on the number of ions, the magnitude of charge on each ion, the ionic mobility (distance travelled by an ion per sec under a potential gradient of one volt/cm).

* Different kinds of Electrolytic Conductance:

1. Specific Conductance: Reciprocal of specific resistance. $\kappa = \frac{1}{\rho} = G \frac{l}{A}$.

($G \rightarrow$ conductance). ($\frac{l}{A} \rightarrow$ cell constant).

Unit - $S m^{-1}$.

κ decreases on dilution as number of ions per cc decreases.

2. Equivalent Conductance (λ_{eq}): Conducting power of all the ions present on one gram equivalent of an electrolyte in a given solution.

$$\lambda_{eq} = \kappa V = \kappa \times \frac{1000}{N} \quad [V \rightarrow \text{volume in cc containing one eq. of electrolyte}]$$

Unit - $S m^2 eq^{-1}$

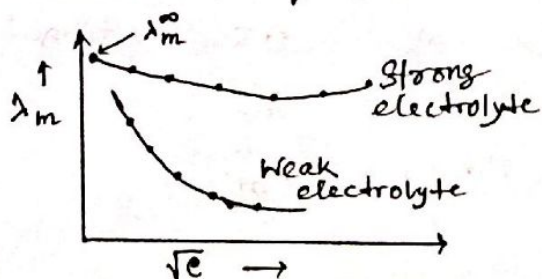
$N \rightarrow$ normality]

λ_{eq} increases with dilution due to large increase in V .

3. Molar Conductance: λ_m - Conducting power of all the ions present in one mole of an electrolyte in a given solution.

$$\lambda_m = \kappa V = \kappa \times \frac{1000}{M} \quad [V \rightarrow \text{volume in cc containing one mole of electrolyte, } M \rightarrow \text{Molarity}]$$

• Variation of λ_m with concentration:



$\lambda_m^\infty \rightarrow$ Molar conductance at infinite dilution (limiting molar conductance)

$c \rightarrow$ Concentration.

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• Debye-Huckel Onsager Equation: $\lambda_m = \lambda_m^\infty - b\sqrt{c}$

[$\lambda_m \rightarrow$ molar conductivity at \sqrt{c} -dilution

$\lambda_m^\infty \rightarrow$ molar conductivity at infinite dilution

$b \rightarrow$ constant that depends upon nature of solvent & temperature, $c \rightarrow$ conc.]

* Kohlrausch's Law: Equivalent conductivity of any electrolyte at infinite dilution is the sum of the limiting ionic conductivities of the cations and anions or molar conductivity of an electrolyte at infinite dilution is the sum of the limiting ionic conductivities of the cations & anions each multiplied by the number of ions present in one formula unit of the electrolyte.

$$\lambda_{eq}^\circ = \lambda_c^\circ + \lambda_a^\circ \quad \Bigg| \quad \lambda_m^\circ \text{ for } A_xB_y = x\lambda_A^\circ + y\lambda_B^\circ$$

• Ionic mobility (u) = $\frac{\text{Ionic velocity}}{\text{Potential Gradient}}$

* Applications of Conductance:

1. Calculation of molar conductivity of weak electrolytes:

$$\lambda_m^\circ \text{CH}_3\text{COOH} = \lambda_{\text{CH}_3\text{COO}^-}^\circ + \lambda_{\text{H}^+}^\circ = (\lambda_{\text{H}^+}^\circ + \lambda_{\text{Cl}^-}^\circ) + (\lambda_{\text{CH}_3\text{COO}^-}^\circ + \lambda_{\text{Na}^+}^\circ) - (\lambda_{\text{Na}^+}^\circ + \lambda_{\text{Cl}^-}^\circ)$$

$$= \lambda_{m\text{HCl}}^\circ + \lambda_{m\text{CH}_3\text{COONa}}^\circ - \lambda_{m\text{NaCl}}^\circ$$

2. Calculation of degree of dissociation:

$$\alpha = \frac{\lambda_m \text{ at concentration } c}{\lambda_m \text{ at infinite dilution}} = \frac{\lambda_m^c}{\lambda_m^\circ}$$

3. Calculation of dissociation constant (K_c)

of Weak Electrolyte: $AB \rightleftharpoons A^+ + B^-$

$$K_c = \frac{c\alpha \cdot c\alpha}{c - c\alpha} = \frac{c\alpha^2}{1 - \alpha}$$

4. Determination of Ionic product of water (K_w):

$$\lambda_{H_2O}^0 = \lambda_{H^+}^0 + \lambda_{OH^-}^0 = K \times \frac{1000}{M}$$

$$\Rightarrow M = [H^+] = [OH^-] = \frac{K \times 1000}{\lambda_{H^+}^0 + \lambda_{OH^-}^0} = 1.01 \times 10^{-7} \text{ g ions/L.}$$

$$K_w = [H^+][OH^-] = 1.02 \times 10^{-14}$$

5. Determination of solubility of a sparingly

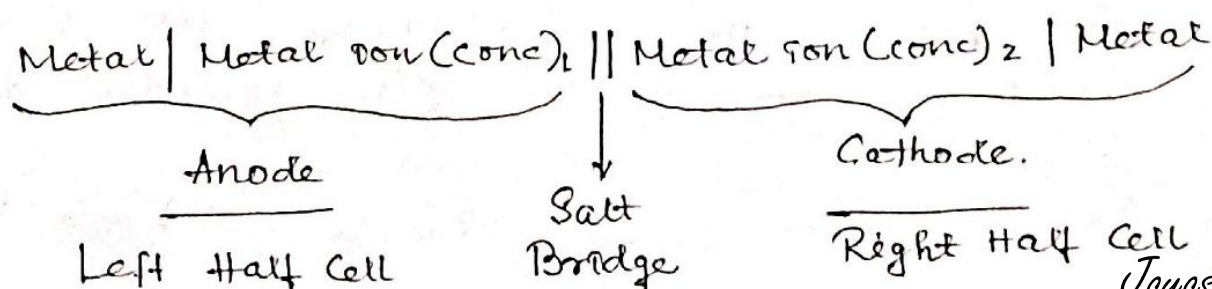
soluble salt: $\lambda_m^0 = K \times \frac{1000}{M} = K \times \frac{1000}{\text{Solubility}}$

$$\text{Solubility} = \frac{K \times 1000}{\lambda_+^0 + \lambda_-^0}$$

* Electrochemical Cells: Cells that convert chemical energy to electrical energy.

* Salt Bridge: i) Prevents the mechanical mixing of solution, ii) Completes electrical circuit, iii) maintains electrical neutrality. iv) prevents development of liquid junction potential.

* Representation of an electrochemical cell:



* **Electrode:** Device in which a metallic (or electronic) conductor is in contact with electrolytic conductor is called electrode.
 1) Electrode behaves as anode if oxidation takes place. 2) Electrode behaves as cathode if reduction takes place.

• **Electrode Signs.**

Electrolytic Cell		Electrochemical Cell	
Anode	Cathode	Anode	Cathode
+	-	-	+
Oxidation	Reduction	Oxidation	Reduction.

* **Electrode potential:** Tendency of an electrode to lose or gain electrons when it is in contact with the solution of its own ions. // If standard conditions ($T = 298\text{ K}$, $\text{conc.} = 1\text{ M}$, $P = 1\text{ atm}$) are used, the electrode potential is called standard electrode potential (SEP). // The tendency to lose electrons is called oxidation potential. // The tendency to gain electrons is called reduction potential. // The SEP of an electrode is determined with respect to standard hydrogen electrode (SHE) whose SEP is taken as 0.

* **EMF:** The electrode potential difference between the two half cells is known as EMF of the cell or cell potential or cell voltage.

$$E_{\text{cell}}^{\circ} = E_{\text{red}}^{\circ}(\text{cathode}) - E_{\text{red}}^{\circ}(\text{anode}) = E_{\text{right}}^{\circ} - E_{\text{left}}^{\circ}$$

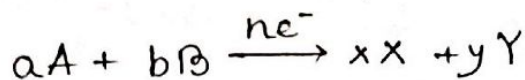
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† The Nernst Equation:

For a reduction reaction, $M^{n+} + ne^- \rightarrow M$

$$\begin{aligned} E_{\text{cell}} &= E_{\text{cell}}^{\circ} - \frac{RT}{nF} \ln \frac{[M(s)]}{[M^{n+}]} \\ &= E_{\text{cell}}^{\circ} - \frac{RT}{nF} \ln \frac{1}{[M^{n+}]} \\ &= E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log \frac{1}{[M^{n+}]} \end{aligned}$$

- Application - Electrode potential of a cell.



$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log \frac{[X]^x [Y]^y}{[A]^a [B]^b}$$

Equilibrium Constant (K_c).

at equilibrium,

$$E_{\text{cell}}^{\circ} = \frac{0.0591}{n} \log K_c$$

Derivation - $\Delta G = \Delta G^{\circ} + RT \ln Q$.

Now, $\Delta G = -nFE$, $\Delta G^{\circ} = -nFE_{\text{cell}}^{\circ}$

$$-nFE_{\text{cell}} = -nFE_{\text{cell}}^{\circ} + RT \ln Q.$$

$$\Rightarrow E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{RT}{nF} \ln \frac{[X]^x [Y]^y}{[A]^a [B]^b}$$

$$\Rightarrow E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log \frac{[X]^x [Y]^y}{[A]^a [B]^b}$$

$$\begin{aligned} [R &\rightarrow 8.31 \text{ J/mol K}, T \rightarrow 298 \text{ K}, \\ F &= 96500 \text{ C}] \end{aligned}$$

$$\dagger \Delta G^\circ = -2.303 RT \log K_c$$

† Applications of EMF:

1. Determination of K_c :

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0591}{n} \log Q$$

at equ^m, $E_{\text{cell}} = 0$, $Q = K_c$.

$$E_{\text{cell}}^\circ = \frac{0.0591}{n} \log K_c$$

$$\Rightarrow \log K_c = \frac{n E_{\text{cell}}^\circ}{0.0591} \Rightarrow K_c = \text{anti log} \left(\frac{n E_{\text{cell}}^\circ}{0.0591} \right)$$

2. Determination of Solubility Product:

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{RT}{nF} \ln \frac{1}{[M^{n+}]}$$

3. Ionisation constant of weak acid or base:

$$[H^+] = \sqrt{CK} \quad \Bigg| \quad E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0591}{n} \log \frac{[X^+]}{[Y^+]}$$

4. pH of a Solution: $E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0591}{n} \log \frac{[X^+]}{[Y^+]}$



5. Thermodynamic Data: $\Delta G = \Delta H + T \left[\frac{d(\Delta G)}{dT} \right]_p$

(Gibbs-Helmholtz Equation)

$$\Rightarrow -nFE_{\text{cell}} = \Delta H - nFT \left[\frac{dE_{\text{cell}}}{dT} \right]_p$$

$$\Rightarrow \Delta H = -nF \left[E_{\text{cell}} - T \left(\frac{dE_{\text{cell}}}{dT} \right)_p \right]$$

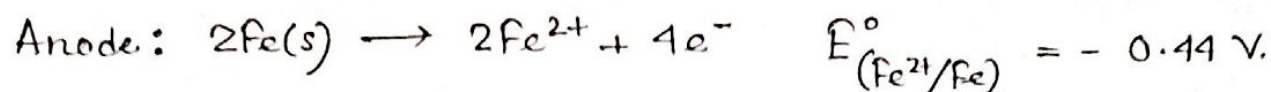
$$\Rightarrow \left(\frac{dE_{\text{cell}}}{dT} \right)_p = \frac{\Delta H}{nFT} + \frac{E_{\text{cell}}}{T}$$

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

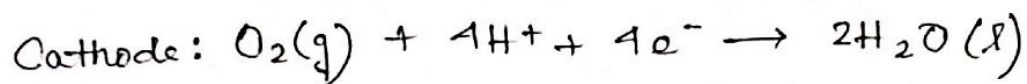
$$\Rightarrow \Delta S = - \left(\frac{d(\Delta G)}{dT} \right)_p = - \left[\frac{d(-nFE_{\text{cell}})}{dT} \right]_p = nF \left(\frac{dE_{\text{cell}}}{dT} \right)_p$$

* Corrosion: (Rusting of iron, tarnishing of silver, development of green coating on Cu).

Rusting of iron occurs in presence of water & air. Corrosion is considered to be an electrochemical phenomenon. At a particular spot of an iron material, oxidation takes place & that spot behaves as anode.

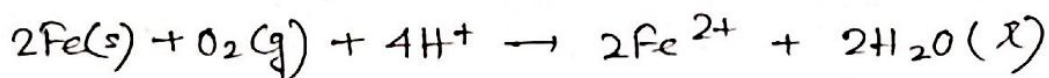


Electrons released at anodic spot move through the metal & go to another spot on the metal & reduce oxygen. This place acts as cathode.



$$E^\circ_{\text{H}^+|\text{O}_2|\text{H}_2\text{O}} = 1.23 \text{ V.}$$

Overall reaction -



$$E^\circ_{\text{cell}} = 1.67 \text{ V.}$$

The ferrous ions are further oxidised by atmospheric oxygen to ferric ions which come out as rust in the form of hydrated ferric oxide ($\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$).