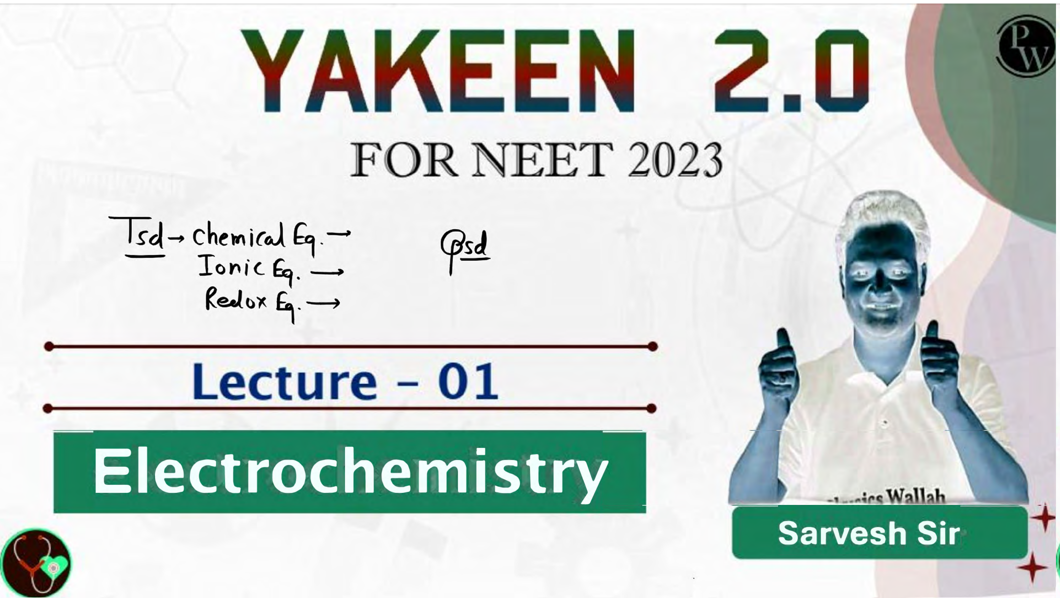


YAKEEN 2.0

FOR NEET 2023

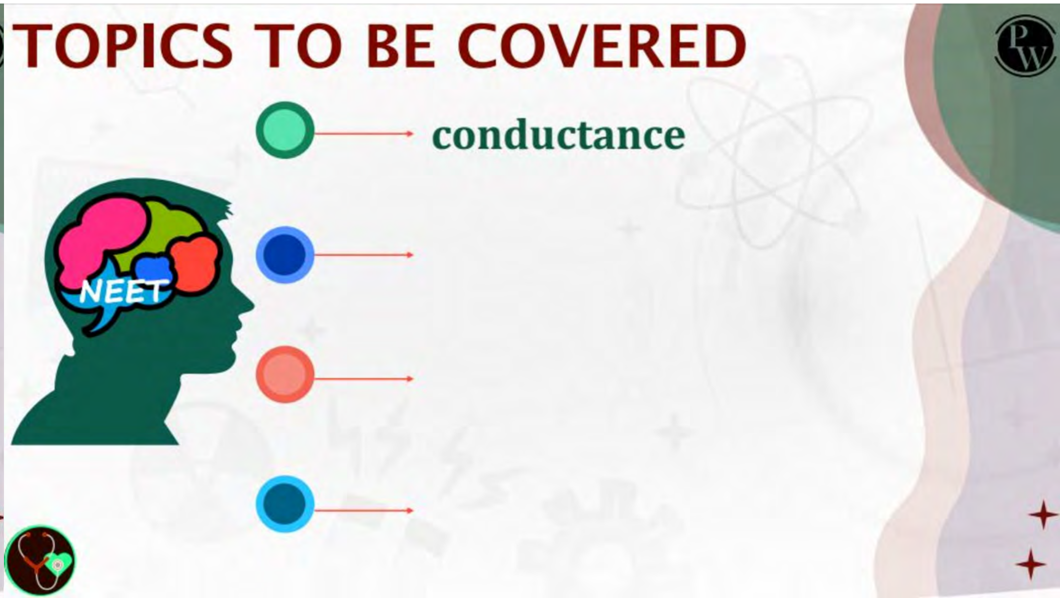
Tsd → chemical Eq. →
Ionic Eq. →
Redox Eq. →

Qsd

Lecture – 01

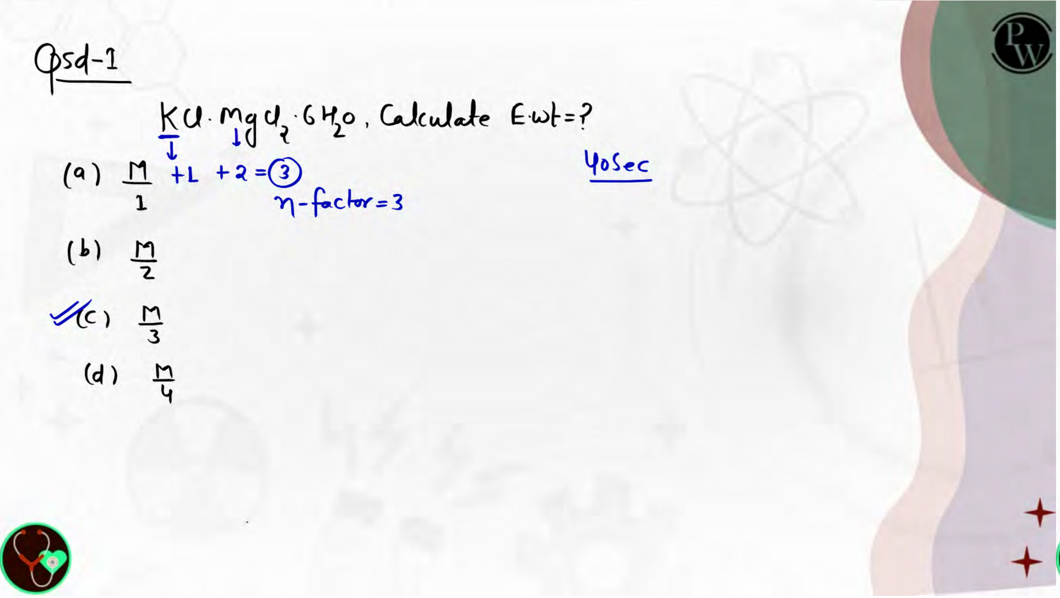
Electrochemistry

Sarvesh Sir



TOPICS TO BE COVERED

- conductance



Qsd-1

$KCl \cdot MgCl_2 \cdot 6H_2O$, Calculate E.wt = ?

40Sec

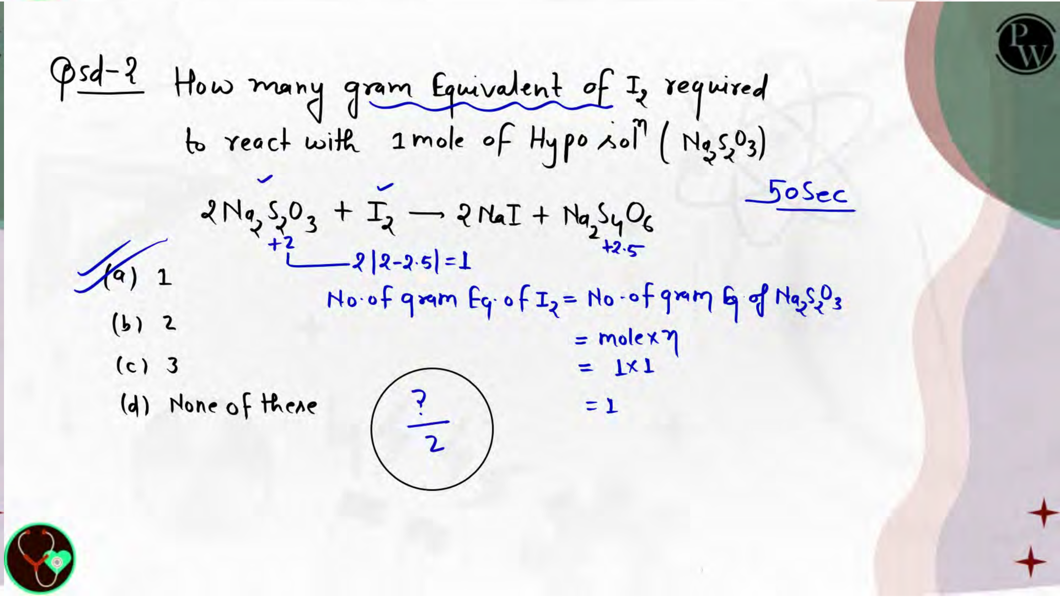
(a) $\frac{M}{1}$ $+1 \; +2 = \boxed{3}$

η-factor = 3

(b) $\frac{M}{2}$

✓(c) $\frac{M}{3}$

(d) $\frac{M}{4}$



Qsd-2 How many gram Equivalent of I_2 required to react with 1 mole of Hypo solⁿ ($Na_2S_2O_3$)

50Sec

$$2Na_2\overset{+2}{S_2}O_3 + I_2 \longrightarrow 2NaI + Na_2\overset{+2.5}{S_4}O_6$$

$2|2-2.5| = 1$

✓(a) 1

(b) 2

(c) 3

(d) None of these

No. of gram Eq. of I_2 = No. of gram Eq of $Na_2S_2O_3$
= mole × η
= 1 × 1
= 1

$\frac{?}{2}$

Electro-chemistry → 2 to 4 Questⁿ

① Conductance

0 - 1 Questⁿ

Cell

1-2 Questⁿ → 100%

Electrolysis + battery

0 - 1 Questⁿ

Conductance

→ $C = \frac{1}{R}$ → mathmatical form

Conductance = $\frac{1}{\text{Resistance}}$

Conductance is equal to $\frac{1}{\text{Resistance}}$ (Reciprocal of Resistance)

" It is measure rate of flow of charge in electrolyte solution "

Unit of Conductance

$\rightarrow C = \frac{1}{R} = \frac{1}{\Omega} = \Omega^{-1}$ or Siemen's ($\underline{S}$)

Effacting factor

CsD = 1

Conductance ∝ No. of ions

Ex → ⓐ 1 mole NaCl solⁿ

✓✓ (b) 2 mole NaCl solⁿ

Arrange Conductance

$NaCl \rightarrow Na^+ + Cl^-$

1 mole NaCl = 2 ion

2 mole — = 4 ion

Ex → Arrange Conductance

ⓐ 1 mole HCl

✓✓ (b) 1 mole H_2SO_4

solⁿ → $\underline{HCl} \rightarrow \underline{H^+ + Cl^-}$

2 ion

$H_2SO_4 \rightarrow \underline{2}H^+ + \underline{S}O_4^{-2}$

1 mole = 3 mole ion

(2) Nature of electrolyte

Csd = 2

Conductance → Strong electrolyte > weak electrolyte

Ex → Arrange Conductance

(a) 1 mole CH_3COOH

✓ (b) 1 mole CH_3COONa

Sol^n → CH_3COOH → Weak Acid

CH_3COONa → salt
↓
100%

Q select the solution which have maximum Conductance.

(a) 1 mole H_3PO_4 → 4 ion

✓ (b) 1 mole H_2SO_4 → 3 ion

(c) 1 mole CH_3COOH

(d) None of these

Sol^n → $\left.\begin{matrix} H_3PO_4 \\ CH_3COOH \end{matrix}\right\} \to$ W.A

Effect of size of ions

Csp = 3

→ Conductance ∝ size of ion (gas ion) ∝ $\frac{1}{\text{Hydrated ion}}$

gas ion → $Na^+ < K^+$

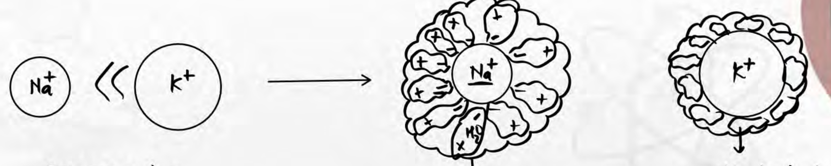
hydrated ion
→ $Na^+(aq) > K^+(aq)$

Ex → $Li^+(aq)$, $Na^+(aq)$, $K^+(aq)$, $Rb^+(aq)$, $Cs^+(aq)$

Conductance → $Cs^+(aq) > Rb^+(aq) > K^+(aq) > Na^+(aq) > Li^+(aq)$

Ex → select the correct ion which have max. Conductance

(a) $Be^{+2}_{(aq)}$ (b) $Mg^{+2}_{(aq)}$ (c) $Ca^{+2}_{(aq)}$ ✓ (d) $Ba^{+2}_{(aq)}$



$F = K\frac{q_1 q_2}{r^2}$

Conductance ∝ size of gas ion
↓
Conductance ∝ $\frac{1}{\text{size of Hydrated ion}}$

⊗ → Conductance ∝ $\frac{1}{\text{size of Hydrated ion}}$

$C_{SD} = 4$

$$\text{Conductance} \propto \frac{1}{\text{Charge of ion}}$$ → dominating rule

Ex → Al^{+3}, Ba^{+2}, K^{+}

$Al^{+3} < Ba^{+2} < K^{+}$ → Conductance

Ex →

Q select Correct order of Conductance.

(a) $Na^{+}(aq) > K^{+}(aq)$

[✓ (b)] $I^{-}(aq) > F^{-}(aq)$

(c) $Mg^{+2}(aq) > Na^{+}(aq)$

(d) $Al^{+3}(aq) > Mg^{+2}(aq)$

Ex → Arrange Conductance

(a) $Na^{+}(aq)$

(b) $Ba^{+2}(aq)$

[✓ (1)] a > b

(2) b > a

(3) a = b

(4) None

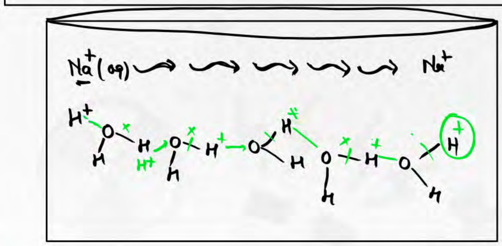
$C_{SD} = 5$

Conductance of H^{+} and OH^{-} is very very high to Grothus Effect

→ $H^{+} >> Na^{+}$

→ $OH^{-} >>> Cl^{-}$

→ strong electrolyte



Ex →

ⓐ 1 mole CH_3COOH → W.A.

(b) 1 mole NaCl → S. Electro-lyte

Conductance → a < b

Ex → ⓐ 1 mole HCl

ⓑ 1 mole NaCl

Conductance → a > b

Q-1 select Correct option for Conductance

Q-1 (a) $SO_4^{-2}(aq) < Cl^{-}(aq)$

(b) $K^{+}(aq) > Na^{+}(aq)$

(c) $K^{+}(aq) < Na^{+}(aq)$

[✓ (d)] 'a' & 'b' Correct

Q-2 → (a) $Na^{+}(aq) > K^{+}(aq) > Rb^{+}(aq)$

(b) $Na^{+}(aq) < K^{+}(aq) < Rb^{+}(aq)$

(c) $H^{+}aq > Na^{+}(aq)$

[✓ (d)] b & c correct

Q-3

(a) $HCl > CH_3COOH$

(b) $NaCl > CH_3COOH$

(c) $H_2SO_4 > HCl$

[✓ (d)] All

$C_{SD} = 6$

- → Conductance ∝ Acidic strength
- → Conductance ∝ basic strength

A.S ⇒ $HCOOH > CH_3COOH$

Ex → ⓐ $HCOOH$, ⓑ CH_3COOH

✓ (a) → a > b
(b) → a < b
(c) → a = b
(d) None

Conductivity or specific Conductance (K = Kappa)

Mathmatical form ⇒ $K = \frac{1}{\rho}$

$\text{Conductivity} = \frac{1}{\text{Resistivity}}$

→ $R \propto \frac{l}{A}$

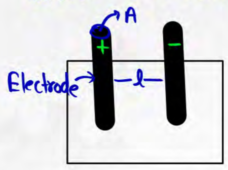
$R = \rho \frac{l}{A}$

$\boxed{\frac{1}{\rho}} = \boxed{\frac{1}{R}} \times \frac{l}{A}$

$K = C \times \frac{l}{A}$

Conductivity = Conductance x cell Constant

where → cell Constant = $\frac{l}{A}$



Unit of Conductivity (K)

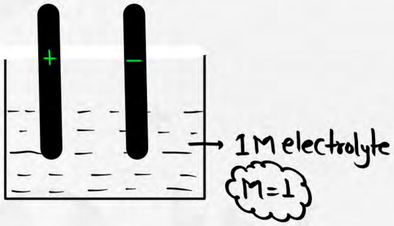
$K = C \times \frac{l}{A}$

$= S \times \frac{Cm}{Cm^2} = S \times Cm^{-1}$ or $S \times m^{-1}$

It is measure rate of flow of charge in $1 cm^3$ electrolyte solⁿ.

Molar Conductance (Λ_m)

→ It is measure rate of flow of charge in 1 Molar electrolyte solution



$$\boxed{\Lambda_m = \frac{K \times 1000}{M}} \rightarrow \Lambda_m \propto \frac{1}{M}$$

K → Conductivity

Λ_m → Molar Conductance

M → Molarity of electrolyte

Unit of λ_m

$$\lambda_m = \frac{K \times 10^3}{M}$$

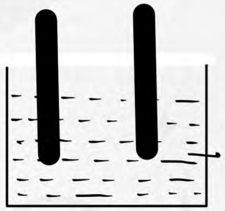
$$= \frac{S \cdot \times Cm^{-1}}{\frac{mole}{Cm^3}}$$

$$= \frac{S \cdot \times Cm^{-1} \times Cm^3}{mole}$$

$$= \underline{S \cdot mole^{-1}\, Cm^2}$$

Equivalent Conductance (λ_e)

→ It is measure rate of flow of charge in 1 Normal electrolyte solution



$$\lambda_e = \frac{K \times 1000}{N}$$

Where → λ_e = Eq. Conductance

N = Normality of electrolyte

K = Conductivity

Unit of λ_e.

$$\lambda_e = \frac{K \times 10^3}{N}$$

$$= \frac{S \times Cm^{-1}}{\frac{gram\ Eq.}{Cm^3}}$$

$$= \frac{S \times Cm^{-1} \times Cm^3}{No. of\ gram\ Eq.} = \boxed{S \times Cm^2 \times gram\ Eq^{-1}}$$

Relation between λ_m and λ_e

$$\lambda_m = \frac{K \times 10^3}{M}$$

$$\lambda_e = \frac{K \times 10^3}{N}$$

$$\frac{\lambda_m}{\lambda_e} = \frac{N}{M} = \frac{\eta\text{-factor} \times \cancel{M}}{\cancel{M}}$$

$$\frac{\lambda_m}{\lambda_e} = \eta\text{-factor}$$

$$\boxed{\lambda_m = \eta\text{-factor} \times \lambda_e}$$

η-factor of ions ⇒ Charge of ions

$Zn^{+2} \rightarrow \eta = 2$, $PO_4^{-3} \rightarrow \eta = 3$

$Al^{+3} \rightarrow \eta = 3$ $SO_4^{-2} \rightarrow \eta = 2$

$Zr^{+4} \rightarrow \eta = 4$ $NO_3^{-} \rightarrow \eta = 1$

$Cu^{+1} \rightarrow \eta = 1$

η-factor of salt – total +ve charge or –ve charge in 1 mole salt

$Na_2CO_3 \Rightarrow \eta = 2$

$Ca_3(PO_4)_2 \rightarrow \eta = 6$

Q Calculate molar Conductance of $Al_2(SO_4)_3$ If Equivalent cond^c of $Al_2(SO_4)_3$ is 10 $S \cdot cm^2\,gram\,Eq^{-1}$.

$2Al^{+3} \Rightarrow \eta = 6$

(a) 10 $S\cdot cm^2\,mole^{-1}$
(b) 30 $S\cdot cm^2\,mole^{-1}$
[✓ (c) 60 $S\cdot cm^2\,mole^{-1}$]
(d) None

$\Lambda_m = \eta \times \Lambda_e$
$= 6 \times 10$
$= 60\ S\cdot cm^2\,mole^{-1}$

Q Equivalent Conductance of Al^{+3} is 10 $S\cdot cm^2\,gram\,Eq^{-1}$. than Calculate molar Conductance of Al^{+3}.

(a) 10 $S\cdot cm^2\,mole^{-1}$
[✓ (b) 30 $S\cdot cm^2\,mole^{-1}$]
(c) 60 $S\cdot cm^2\,mole^{-1}$
(d) None

$\Lambda_m = \eta \times \Lambda_e$
$= 3 \times 10$
$= 30\ S\cdot cm^2\,mole^{-1}$

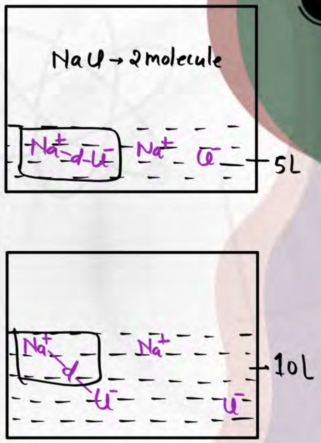
Effect of dilution

dilution ⇒ addition of solvent only (water)

Concentration → Molarity = $\frac{Mole}{Volume}$ ↓

Normality = $\frac{gram\ Eq.}{Volume}$

→ dilution → Concentrat^n (M, N) ↓ → decreases



→ distance b/w ions – increases
→ Collision b/w ions → probability → Kam
} → motion of ions – increases

→ Conductance ⇒ $C = \frac{1}{R}$

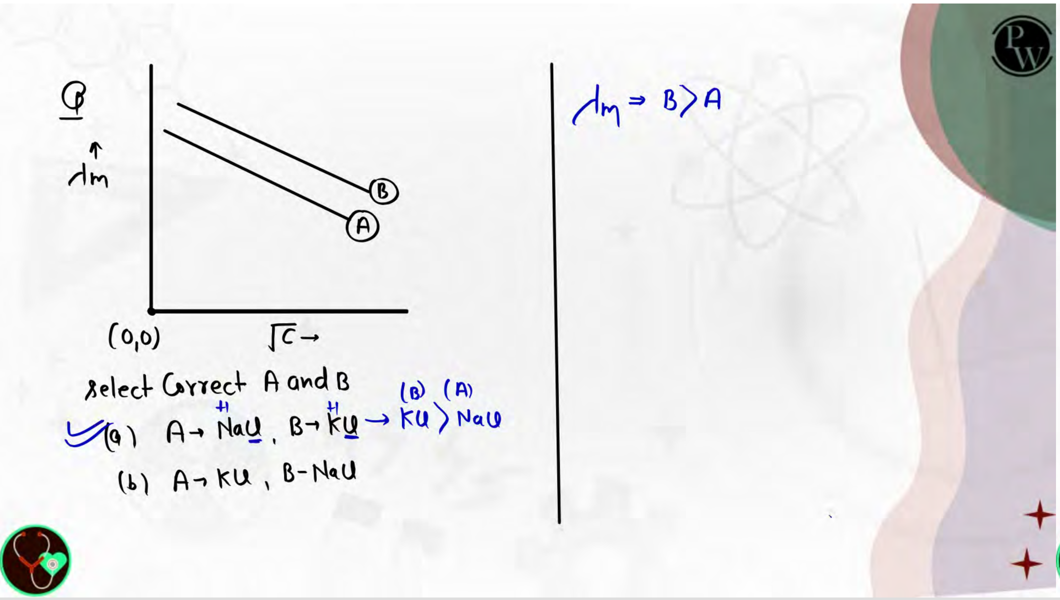
→ molar conductance (Λ_m) ⇒ $\Lambda_m = \frac{K \times 10^3}{M}$

→ Equivalent Conductance (Λ_e) ⇒ $\Lambda_e = \frac{K \times 10^3}{N}$

} → increases on dilution

→ Conductivity or specific Conductance } → decreases on dilution

Reason ⇒ No. of ions per cm^3 decreases
So, Conductivity decreases



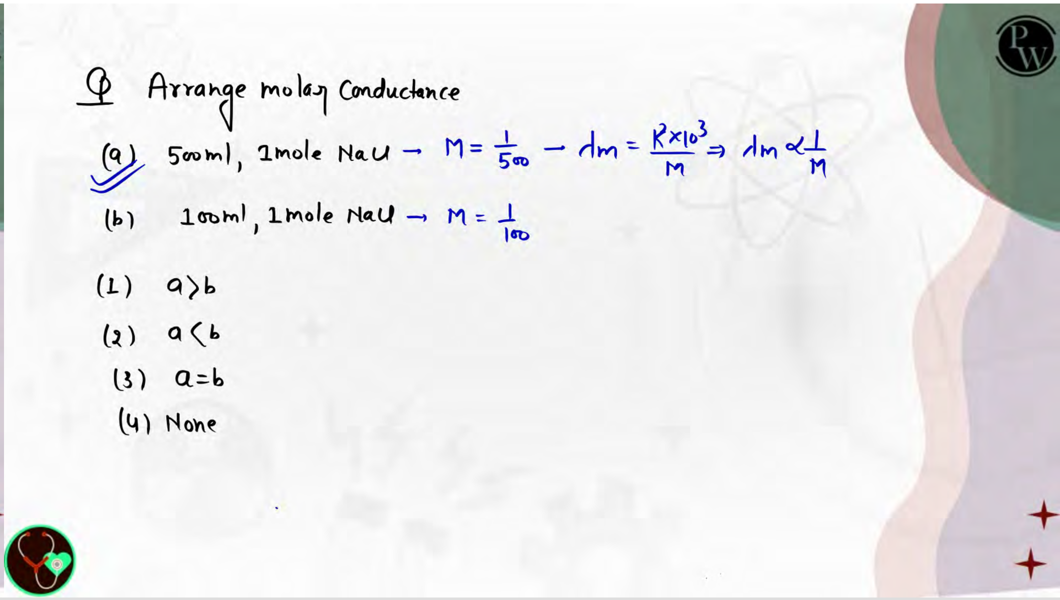
Q

select Correct A and B

✓ (a) $A \rightarrow \overset{H}{Na}Cl$, $B \rightarrow \overset{H}{K}Cl \rightarrow \overset{(B)}{KCl} > \overset{(A)}{NaCl}$

(b) $A \rightarrow KCl$, $B - NaCl$

$\lambda_m \Rightarrow B > A$



Q Arrange molar Conductance

✓ (a) 500ml, 1mole NaCl → $M = \frac{1}{500}$ → $\lambda_m = \frac{K \times 10^3}{M} \Rightarrow \lambda_m \propto \frac{1}{M}$

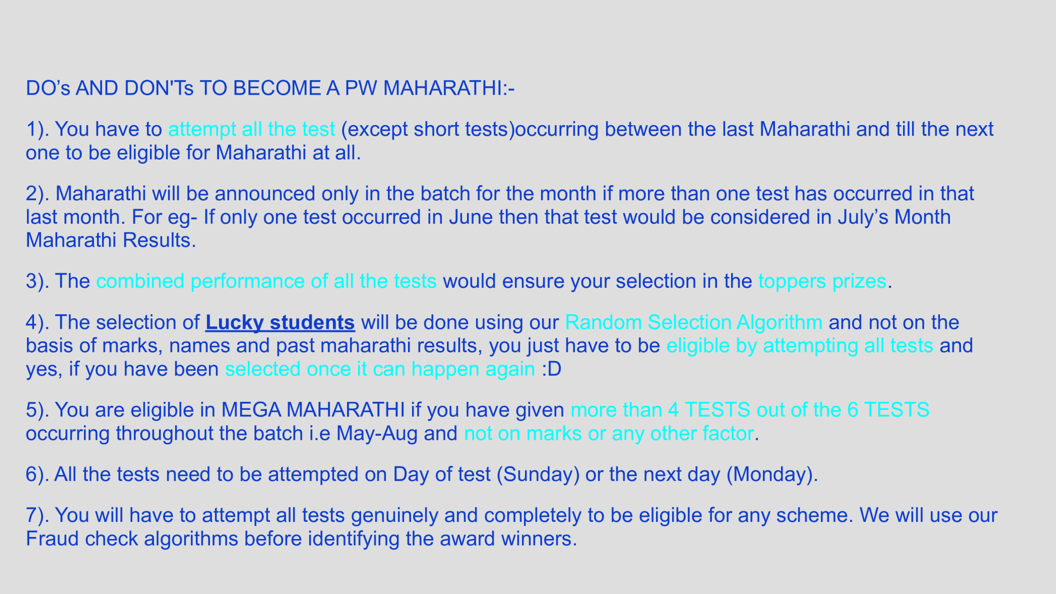
(b) 100ml, 1mole NaCl → $M = \frac{1}{100}$

(1) a > b

(2) a < b

(3) a = b

(4) None



DO’s AND DON'Ts TO BECOME A PW MAHARATHI:-

1). You have to attempt all the test (except short tests)occurring between the last Maharathi and till the next one to be eligible for Maharathi at all.

2). Maharathi will be announced only in the batch for the month if more than one test has occurred in that last month. For eg- If only one test occurred in June then that test would be considered in July’s Month Maharathi Results.

3). The combined performance of all the tests would ensure your selection in the toppers prizes.

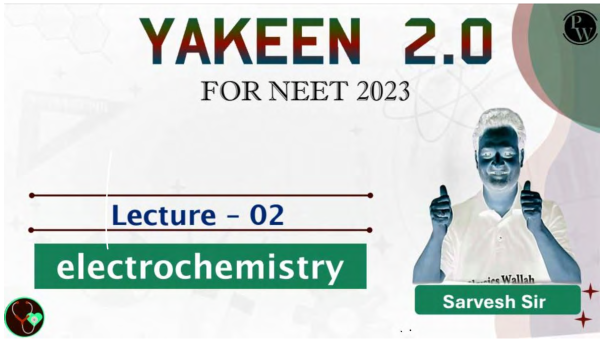
4). The selection of **Lucky students** will be done using our Random Selection Algorithm and not on the basis of marks, names and past maharathi results, you just have to be eligible by attempting all tests and yes, if you have been selected once it can happen again :D

5). You are eligible in MEGA MAHARATHI if you have given more than 4 TESTS out of the 6 TESTS occurring throughout the batch i.e May-Aug and not on marks or any other factor.

6). All the tests need to be attempted on Day of test (Sunday) or the next day (Monday).

7). You will have to attempt all tests genuinely and completely to be eligible for any scheme. We will use our Fraud check algorithms before identifying the award winners.

THANK YOU



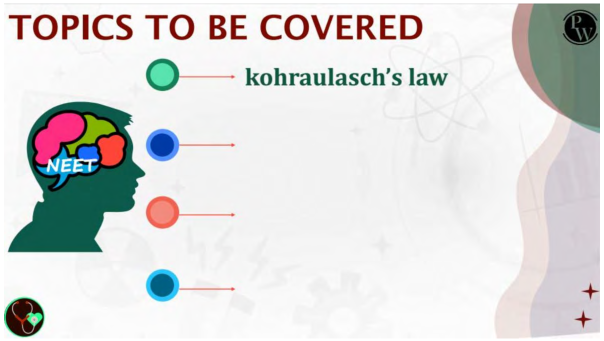
YAKEEN 2.0

FOR NEET 2023

Lecture – 02

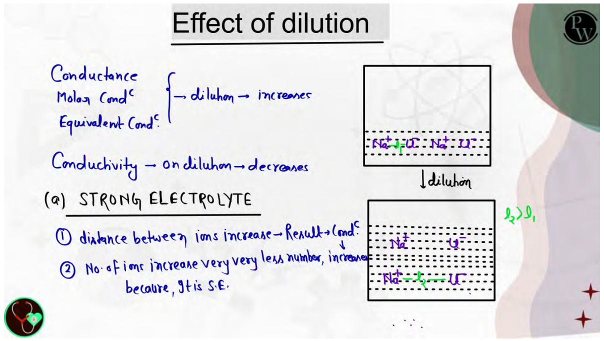
electrochemistry

Sarvesh Sir



TOPICS TO BE COVERED

- kohraulasch’s law



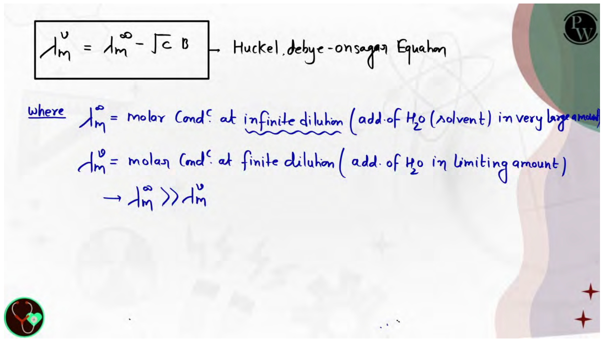
Effect of dilution

Conductance
Molar Cond^c
Equivalent Cond^c } → dilution → increases

Conductivity → on dilution → decreases

(a) STRONG ELECTROLYTE

① distance between ions increase → Result → Cond^c ↓ increases
② No. of ions increase very very less number, because, It is S.E.

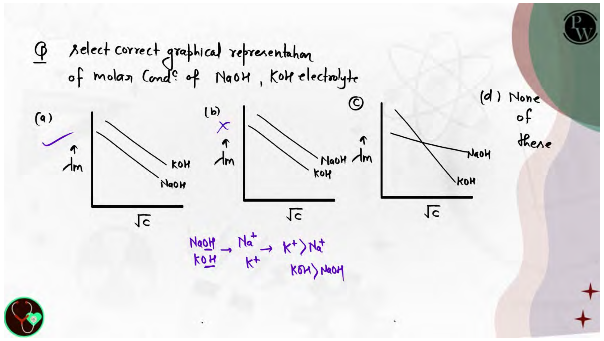


$$\boxed{\Lambda_m^{v} = \Lambda_m^{\infty} - \sqrt{c}\,B}$$ → Huckel, debye-onsagar Equation

where Λ_m^{∞} = molar Cond^c at infinite dilution (add. of H_2O (solvent) in very large amount)

Λ_m^{v} = molar Cond^c at finite dilution (add. of H_2O in limiting amount)

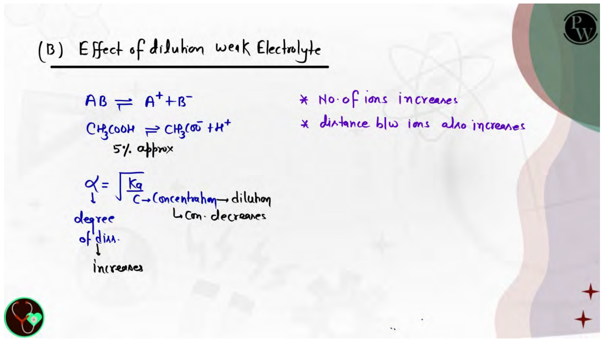
→ $\Lambda_m^{\infty} >> \Lambda_m^{v}$



Q. select correct graphical representation of molar Cond^c of NaOH, KOH electrolyte

(a) ✓
(b) ✗
(c)
(d) None of these

$NaOH \rightarrow Na^+ \rightarrow K^+ > Na^+$
$KOH \quad K^+$
$KOH > NaOH$



(B) Effect of dilution weak Electrolyte

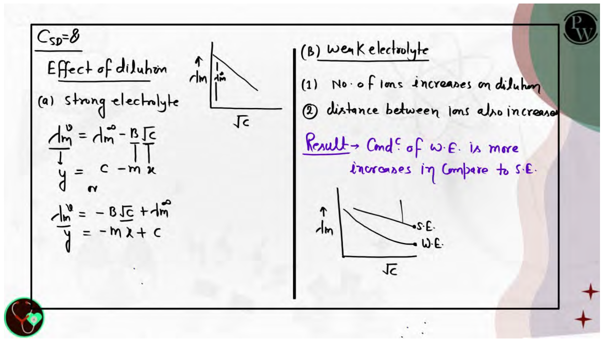
$AB \rightleftharpoons A^+ + B^-$

$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$
5% approx

$\alpha = \sqrt{\frac{Ka}{C}}$ C → Concentration → dilution
└ Con. decreases

α ↓ degree of diss. ↓ increases

* No. of ions increases
* distance b/w ions also increases



$C_{SD} = 8$

Effect of dilution

(a) strong electrolyte

$$\Lambda_m^c = \Lambda_m^\infty - B\sqrt{c}$$
$$y = c - m x$$

or

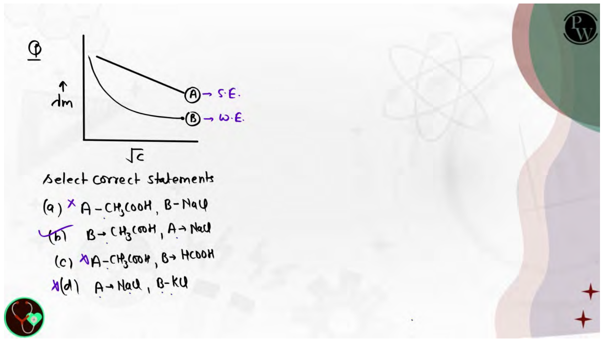
$$\Lambda_m^c = -B\sqrt{c} + \Lambda_m^\infty$$
$$y = -m x + c$$

(B) weak electrolyte

(1) No. of ions increases on dilution

② distance between ions also increase

Result → Cond^c of W.E. is more increases in compare to S.E.



Q

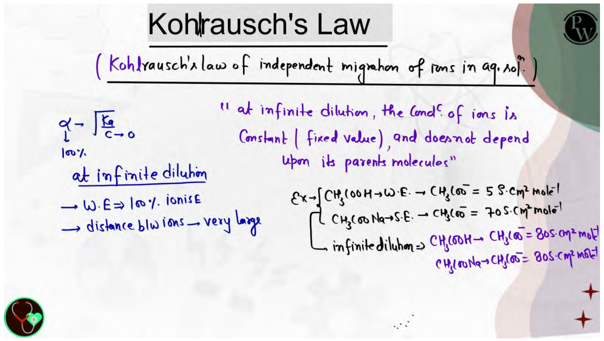
Select correct statements

(a) ✗ A – CH_3COOH, B – NaCl

✓ (b) B → CH_3COOH, A → NaCl

(c) ✗ A – CH_3COOH, B → HCOOH

✗ (d) A → NaCl, B – KCl



Kohlrausch's Law

(Kohlrausch's law of independent migration of ions in aq. sol^n.)

$\alpha \rightarrow \sqrt{\frac{K_a}{c}}$, $c \rightarrow 0$

↓

100%

at infinite dilution

→ W.E ⇒ 100% ionise

→ distance b/w ions → very large

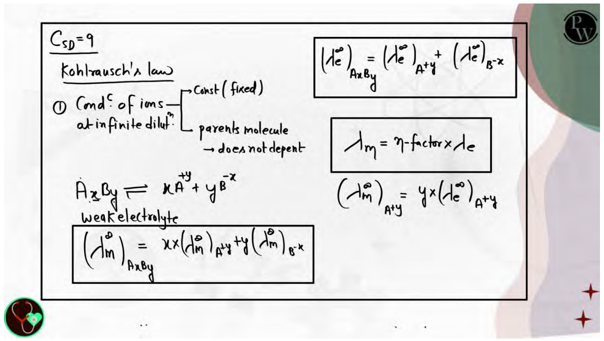
" at infinite dilution, the Cond^c of ions is Constant (fixed value), and does not depend upon its parents molecules"

Ex → CH_3COOH → W.E. → CH_3COO^- = 5 S·cm² mole⁻¹

CH_3COONa → S.E. → CH_3COO^- = 70 S·cm² mole⁻¹

infinite dilution ⇒ CH_3COOH → CH_3COO^- = 80 S·cm² mole⁻¹

CH_3COONa → CH_3COO^- = 80 S·cm² mole⁻¹



$C_{SD} = 9$

Kohlrausch's law

① Cond^c of ions at infinite dilut^n → Const (fixed)

→ parents molecule → does not depent

$$A_xB_y \rightleftharpoons xA^{+y} + yB^{-x}$$

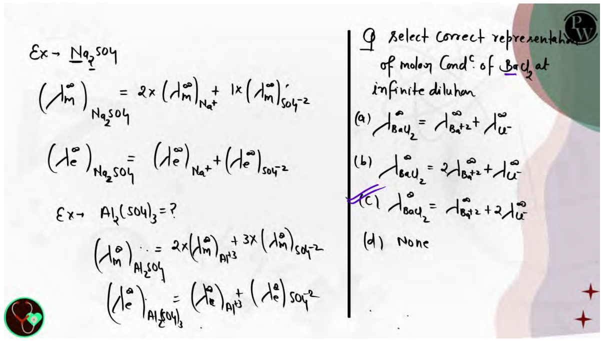
weak electrolyte

$$\boxed{(\Lambda_m^\infty)_{A_xB_y} = x \times (\Lambda_m^\infty)_{A^{+y}} + y(\Lambda_m^\infty)_{B^{-x}}}$$

$$\boxed{(\Lambda_e^\infty)_{A_xB_y} = (\Lambda_e^\infty)_{A^{+y}} + (\Lambda_e^\infty)_{B^{-x}}}$$

$$\boxed{\Lambda_m = \eta\text{-factor} \times \Lambda_e}$$

$$(\Lambda_m^\infty)_{A^{+y}} = y \times (\Lambda_e^\infty)_{A^{+y}}$$



Ex → Na_2SO_4

$$(\lambda_m^\infty)_{Na_2SO_4} = 2\times(\lambda_m^\infty)_{Na^+} + 1\times(\lambda_m^\infty)_{SO_4^{-2}}$$

$$(\lambda_e^\infty)_{Na_2SO_4} = (\lambda_e^\infty)_{Na^+} + (\lambda_e^\infty)_{SO_4^{-2}}$$

Ex → $Al_2(SO_4)_3 = ?$

$$(\lambda_m^\infty)_{Al_2SO_4} = 2\times(\lambda_m^\infty)_{Al^{+3}} + 3\times(\lambda_m^\infty)_{SO_4^{-2}}$$

$$(\lambda_e^\infty)_{Al_2(SO_4)_3} = (\lambda_e^\infty)_{Al^{+3}} + (\lambda_e^\infty)_{SO_4^{-2}}$$

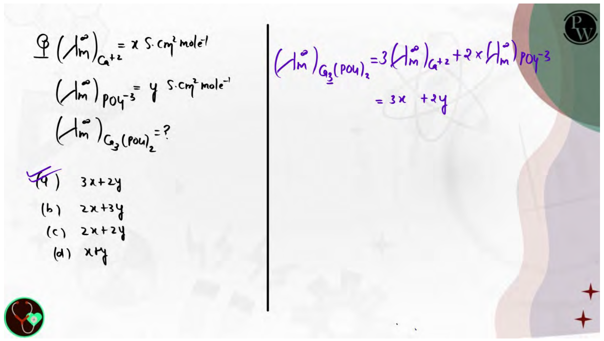
Q Select correct representation of molar Cond^c of $BaCl_2$ at infinite dilution

(a) $\lambda^\infty_{BaCl_2} = \lambda^\infty_{Ba^{+2}} + \lambda^\infty_{Cl^-}$

(b) $\lambda^\infty_{BaCl_2} = 2\lambda^\infty_{Ba^{+2}} + \lambda^\infty_{Cl^-}$

✓ (c) $\lambda^\infty_{BaCl_2} = \lambda^\infty_{Ba^{+2}} + 2\lambda^\infty_{Cl^-}$

(d) None



Q $(\lambda_m^\infty)_{Ca^{+2}} = x \text{ S}\cdot\text{cm}^2\,\text{mole}^{-1}$

$(\lambda_m^\infty)_{PO_4^{-3}} = y \text{ S}\cdot\text{cm}^2\,\text{mole}^{-1}$

$(\lambda_m^\infty)_{Ca_3(PO_4)_2} = ?$

✓ (a) $3x+2y$

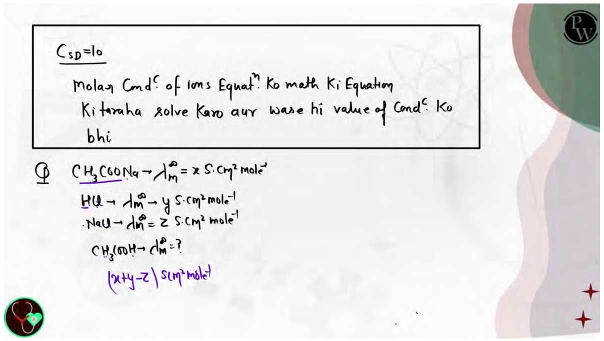
(b) $2x+3y$

(c) $2x+2y$

(d) $x+y$

$$(\lambda_m^\infty)_{Ca_3(PO_4)_2} = 3(\lambda_m^\infty)_{Ca^{+2}} + 2\times(\lambda_m^\infty)_{PO_4^{-3}}$$

$$= 3x + 2y$$



$C_{SD} = 10$

Molar Cond^c of ions Equat^n Ko math Ki Equation Ki taraha solve Karo aur wase hi value of Cond^c Ko bhi

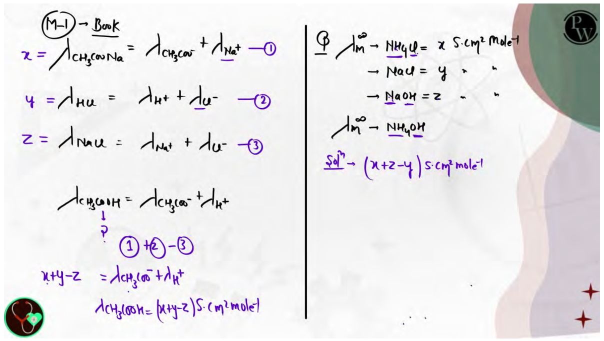
Q $CH_3COONa \rightarrow \lambda_m^\infty = x \text{ S}\cdot\text{cm}^2\,\text{mole}^{-1}$

$HCl \rightarrow \lambda_m^\infty \rightarrow y \text{ S}\cdot\text{cm}^2\,\text{mole}^{-1}$

$NaCl \rightarrow \lambda_m^\infty = z \text{ S}\cdot\text{cm}^2\,\text{mole}^{-1}$

$CH_3COOH \rightarrow \lambda_m^\infty = ?$

$(x+y-z) \text{ Scm}^2\,\text{mole}^{-1}$



Ⓜ-1 → Book

$x = \lambda_{CH_3COONa} = \lambda_{CH_3COO^-} + \lambda_{Na^+}$ —①

$y = \lambda_{HCl} = \lambda_{H^+} + \lambda_{Cl^-}$ —②

$z = \lambda_{NaCl} = \lambda_{Na^+} + \lambda_{Cl^-}$ —③

$$\lambda_{CH_3COOH} = \lambda_{CH_3COO^-} + \lambda_{H^+}$$

↓

?

① + ② − ③

$x+y-z = \lambda_{CH_3COO^-} + \lambda_{H^+}$

$\lambda_{CH_3COOH} = (x+y-z) \text{ S}\cdot\text{cm}^2\,\text{mole}^{-1}$

Q $\lambda_m^\infty \rightarrow NH_4Cl = x \text{ S}\cdot\text{cm}^2\,\text{mole}^{-1}$

→ $NaCl = y$ " "

→ $NaOH = z$ " "

$\lambda_m^\infty \rightarrow NH_4OH$

Sol^n → $(x+z-y) \text{ S}\cdot\text{cm}^2\,\text{mole}^{-1}$

Q $H_2SO_4 \rightarrow \Lambda_m^\infty = x$

$BaCl_2 = y$

$HCl = z$

$\Lambda^{\infty\,m}_{BaSO_4} = ?$

$= (y + x - 2z)\ S\,cm^2\,mole^{-1}$

(a) $x+y-z$

(b) $x+z-y$

(c) $x+y-2z$

(d) None

Application

(a) degree of dissociation and Eq. Constant

$CH_3COOH \begin{cases} \rightarrow \Lambda_m^c \\ \rightarrow \Lambda_m^\infty \end{cases}$

$$\text{degree of dissociation } (\alpha) = \frac{\text{molar cond}^c\text{. at finite dilution } (\Lambda_m^c)}{\text{molar cond}^c\text{. at infinite dilution } (\Lambda_m^\infty)} \text{ or } \frac{\Lambda_e^c}{\Lambda_e^\infty}$$

$AB \rightleftharpoons A^+ + B^-$

$C \quad 0 \quad 0 \quad t=0$

$C(1-\alpha) \quad C\alpha \quad C\alpha$ at Eq.

$$K_a = \frac{\cancel{C}\alpha \times C\alpha}{\cancel{C}(1-\alpha)}$$

$$K_a = \frac{C\alpha^2}{1-\alpha} = \frac{C\left(\frac{\Lambda_m^c}{\Lambda_m^\infty}\right)^2}{\left(1-\frac{\Lambda_m^c}{\Lambda_m^\infty}\right)}$$

$$K_a = \frac{C\times(\Lambda_m^c)^2\times\cancel{(\Lambda_m^\infty)}}{(\Lambda_m^\infty)^2(\Lambda_m^\infty-\Lambda_m^c)}$$

$$\boxed{K_a = \frac{C\times(\Lambda_m^c)^2}{\Lambda_m^\infty(\Lambda_m^\infty-\Lambda_m^c)}}$$

(b) Solubility of sparingly soluble salt

$$\Lambda_m = \frac{K\times1000}{M}$$

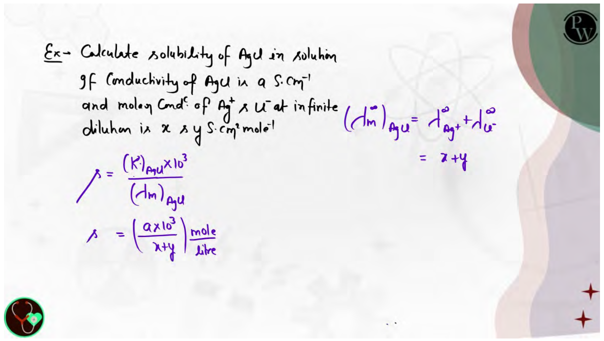
M → Con. of electrolyte (soluble in sol^n.)

$$\boxed{\Lambda_m = \frac{K\times1000}{s}}$$

s (molar solubility = $\frac{\text{mole}}{\text{litre}}$)

s = solubility of salt ($\frac{\text{mole}}{\text{litre}}$)

$$\boxed{s = \frac{K\times10^3}{(\Lambda_m)_{salt}}}$$



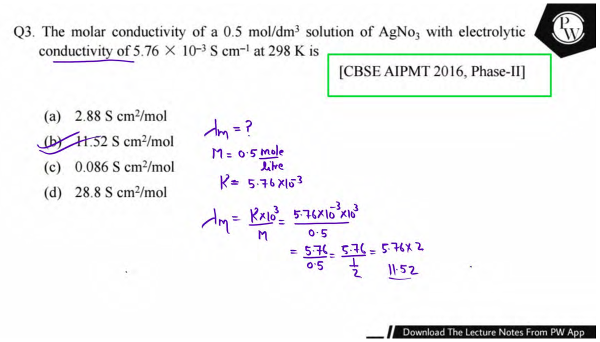
Ex→ Calculate solubility of AgCl in solution
If Conductivity of AgCl is a S·cm⁻¹
and molar Cond^c of Ag^+ & Cl^- at infinite
dilution is x & y S·cm² mole⁻¹

$(\Lambda_m^\infty)_{AgCl} = \Lambda^\infty_{Ag^+} + \Lambda^\infty_{Cl^-}$

$= x + y$

$s = \frac{(K)_{AgCl} \times 10^3}{(\Lambda_m)_{AgCl}}$

$s = \left(\frac{a \times 10^3}{x+y}\right)\frac{mole}{litre}$



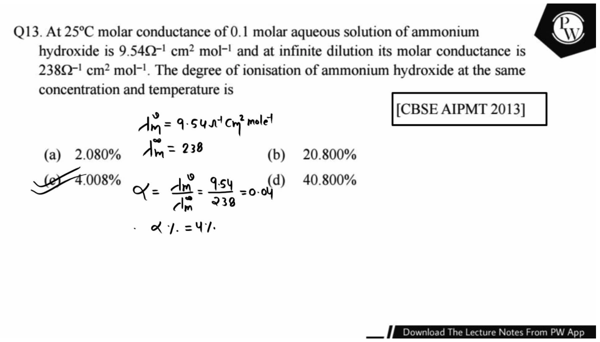
Q3. The molar conductivity of a 0.5 mol/dm³ solution of $AgNo_3$ with electrolytic conductivity of 5.76×10^{-3} S cm⁻¹ at 298 K is

[CBSE AIPMT 2016, Phase-II]

(a) 2.88 S cm²/mol
(b) 11.52 S cm²/mol
(c) 0.086 S cm²/mol
(d) 28.8 S cm²/mol

$\Lambda_m = ?$
$M = 0.5 \frac{mole}{litre}$
$K = 5.76 \times 10^{-3}$

$\Lambda_m = \frac{K \times 10^3}{M} = \frac{5.76 \times 10^{-3} \times 10^3}{0.5}$
$= \frac{5.76}{0.5} = \frac{5.76}{\frac{1}{2}} = 5.76 \times 2$
11.52



Q13. At 25°C molar conductance of 0.1 molar aqueous solution of ammonium hydroxide is 9.54Ω⁻¹ cm² mol⁻¹ and at infinite dilution its molar conductance is 238Ω⁻¹ cm² mol⁻¹. The degree of ionisation of ammonium hydroxide at the same concentration and temperature is

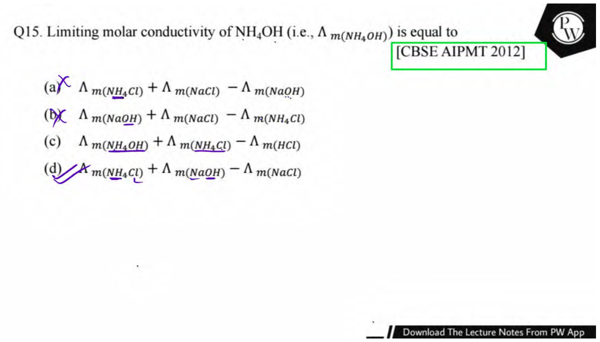
[CBSE AIPMT 2013]

$\Lambda_m^c = 9.54\,\Omega^{-1}\,cm^2\,mole^{-1}$
$\Lambda_m^\infty = 238$

(a) 2.080% (b) 20.800%
(c) 4.008% (d) 40.800%

$\alpha = \frac{\Lambda_m^c}{\Lambda_m^\infty} = \frac{9.54}{238} = 0.04$

$\alpha\% = 4\%$



Q15. Limiting molar conductivity of NH_4OH (i.e., $\Lambda_{m(NH_4OH)}$) is equal to

[CBSE AIPMT 2012]

(a) $\Lambda_{m(NH_4Cl)} + \Lambda_{m(NaCl)} - \Lambda_{m(NaOH)}$
(b) $\Lambda_{m(NaOH)} + \Lambda_{m(NaCl)} - \Lambda_{m(NH_4Cl)}$
(c) $\Lambda_{m(NH_4OH)} + \Lambda_{m(NH_4Cl)} - \Lambda_{m(HCl)}$
(d) $\Lambda_{m(NH_4Cl)} + \Lambda_{m(NaOH)} - \Lambda_{m(NaCl)}$

Q43. Equivalent conductance of NaCl, HCl and C_2H_5COONa at infinite dilution are 126.45, 426.16 and 91 Ω^{-1} cm^2, respectively. The equivalent conductance of C_2H_5COOH is

propanoic acid $91+426.16-126.45$

[CBSE AIPMT 1997]

(a) 201.28 Ω^{-1} cm^2
[✓ (b) 390.71 Ω^{-1} cm^2]
(c) 698.28 Ω^{-1} cm^2
(d) 540.48 Ω^{-1} cm^2

Q40. The specific conductance (κ) of a 0.1 N KCl solution at 23ºC is 0.012 Ω^{-1} cm^{-1}. The resistance of cell containing the solution at the same temperature was found to be 55Ω. The cell constant will be

[CBSE AIPMT 1999]

$C = \frac{1}{R} = \frac{1}{55}$

(a) 0.142 cm^{-1}
[✓ (b) 0.66 cm^{-1}]
(c) 0.918 cm^{-1}
(d) 1.12 cm^{-1}

$\kappa = C \times \frac{l}{A}$ 6

$\frac{l}{A} = \frac{\kappa}{C} = \frac{0.012}{\frac{1}{55}} = 0.012\times55$

$= \underline{.660}$

Q20. An increase in equivalent conductance of a strong electrolyte with dilution is mainly due to

[CBSE AIPMT 2010]

(a) increase in ionic mobility of ions
(b) 100% ionisation of electrolyte at normal dilution
[✓ (c) increase in both, i.e., number of ions and ionic mobility of ions]
(d) increase in number of ions

Q39. The equivalent conductances of Ba^{2+} and Cl^- are 127 and 76Ω^{-1} cm^{-1} eq^{-1} respectively at infinite dilution. The equivalent conductance of $BaCl_2$ at infinite dilution will be

[CBSE AIPMT 2000]

$\lambda_{BaCl_2} = \lambda_{Ba^{+2}} + \lambda_{Cl^-}$

$= 127+76 = \underline{203}$

(a) 139.52
[✓ (b) 203]
(c) 279
(d) 101.5

$(\lambda_m^\infty)_{BaCl_2} = ? = (\lambda_m^\infty)_{Ba^{+2}} + 2(\lambda_m^\infty)_{Cl^-}$

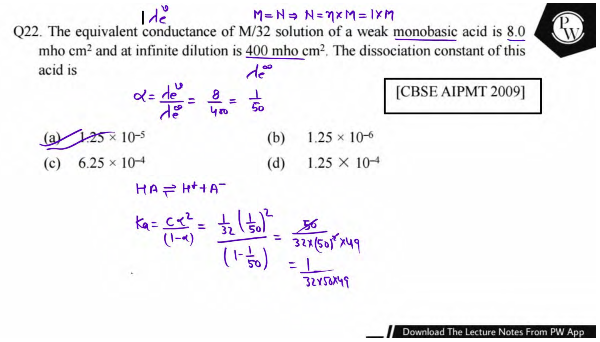
$= 254+76$

$=$

$\lambda_m = n \times \lambda_e$

$(\lambda_m^\infty)_{Ba^{+2}} = 2\times127 = 254$

$(\lambda_m^\infty)_{Cl^-} = 1\times76 = 76$



λ_e^v $\quad M = N \Rightarrow N = n \times M = 1 \times M$

Q22. The equivalent conductance of M/32 solution of a weak monobasic acid is 8.0 mho cm^2 and at infinite dilution is 400 mho cm^2. The dissociation constant of this acid is

λ_e^∞

[CBSE AIPMT 2009]

$$\alpha = \frac{\lambda_e^v}{\lambda_e^\infty} = \frac{8}{400} = \frac{1}{50}$$

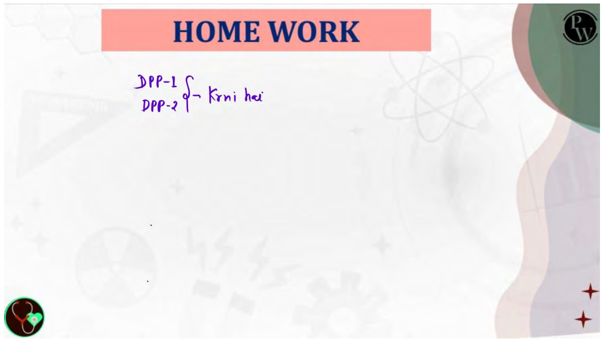
(a) 1.25×10^{-5} (b) 1.25×10^{-6}
(c) 6.25×10^{-4} (d) 1.25×10^{-4}

$$HA \rightleftharpoons H^+ + A^-$$

$$K_a = \frac{C\alpha^2}{(1-\alpha)} = \frac{\frac{1}{32}\left(\frac{1}{50}\right)^2}{\left(1-\frac{1}{50}\right)} = \frac{\cancel{50}}{32 \times (50)^{\cancel{2}} \times 49}$$

$$= \frac{1}{32 \times 50 \times 49}$$

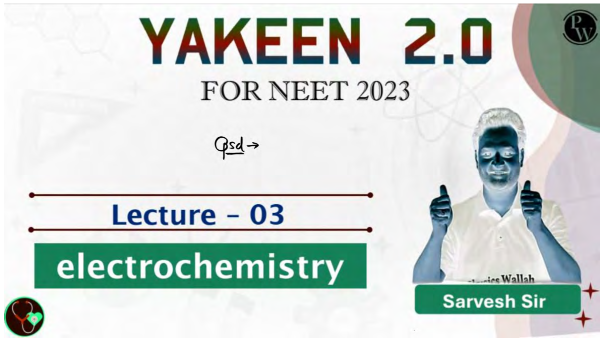
Download The Lecture Notes From PW App



HOME WORK

$$\left.\begin{matrix}\text{DPP-1}\\ \text{DPP-2}\end{matrix}\right\} \rightarrow \text{Krni hai}$$

THANK YOU



YAKEEN 2.0

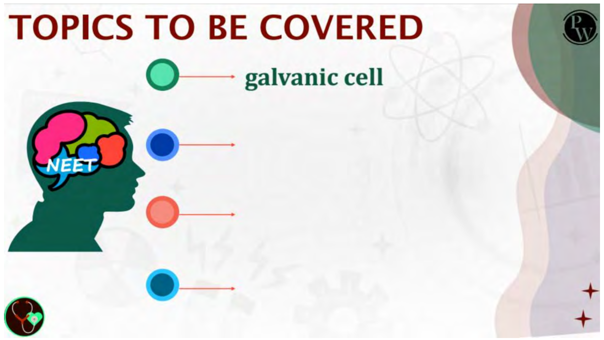
FOR NEET 2023

Qsd →

Lecture – 03

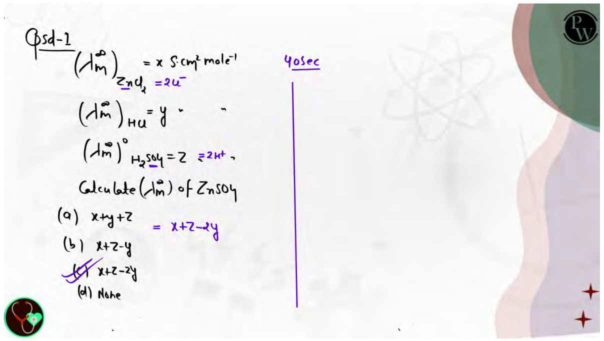
electrochemistry

Sarvesh Sir



TOPICS TO BE COVERED

- galvanic cell



Qsd-1

$(\Lambda_m^\infty)_{ZnCl_2} = x$ S·cm² mole⁻¹ $= 2Cl^-$

$(\Lambda_m^\infty)_{HCl} = y$ " "

$(\Lambda_m^\infty)^o_{H_2SO_4} = z$ $= 2H^+$

Calculate (Λ_m^∞) of $ZnSO_4$

(a) $x+y+z$

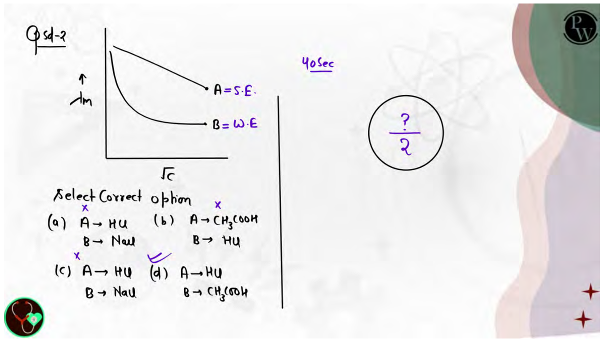
(b) $x+z-y$

[✓ (c)] $x+z-2y$

(d) None

$= x+z-2y$

40sec



Qsd-2

Select Correct option

(a) A → HCl
B → NaCl

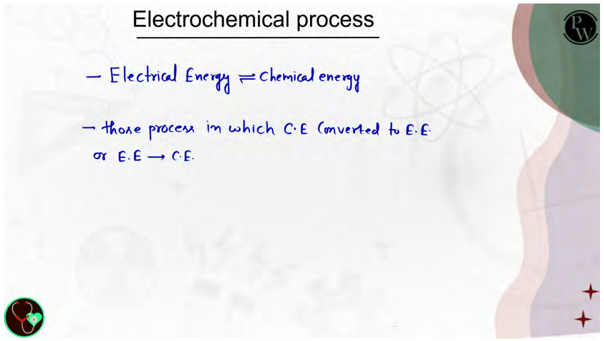
(b) A → CH_3COOH
B → HCl

(c) A → HCl
B → NaCl

[✓ (d)] A → HCl
B → CH_3COOH

40sec

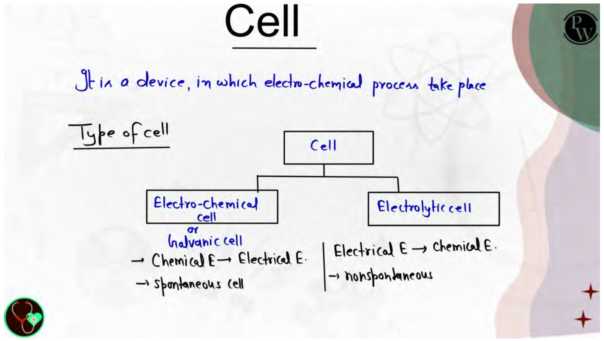
$\frac{?}{2}$



Electrochemical process

– Electrical Energy ⇌ Chemical energy

→ those process in which C.E Converted to E.E.
or E.E ⟶ C.E.

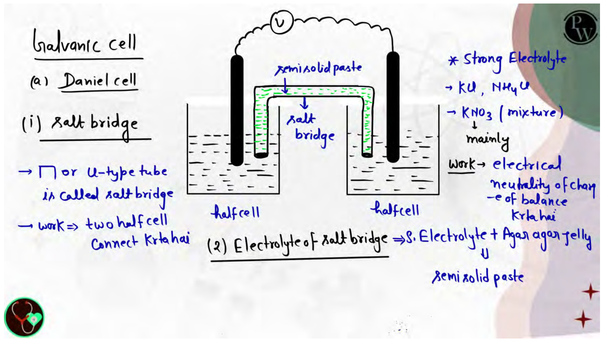


Cell

It is a device, in which electro-chemical process take place

Type of cell

- Cell
  - Electro-chemical cell or Galvanic cell
    - → Chemical E → Electrical E.
    - → spontaneous cell
  - Electrolytic cell
    - Electrical E → chemical E.
    - → nonspontaneous



Galvanic cell

(a) Daniel cell

(i) salt bridge

→ ⊓ or U-type tube is called salt bridge

→ work ⇒ two half cell Connect Krta hai

(2) Electrolyte of salt bridge ⇒ S. Electrolyte + Agar agar jelly
⇓
semisolid paste

* Strong Electrolyte
→ KCl, NH_4Cl
→ KNO_3 (mixture)
↓
mainly

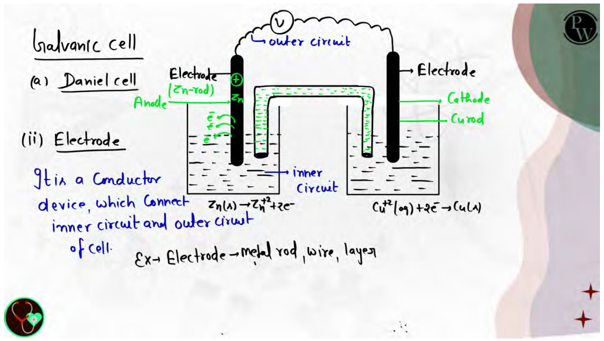
work → electrical neutrality of charge of balance Krta hai

→ ioni mobility (speed of ion) of cation = ionic mobility of Anion

$$(\text{speed})_{\text{Cation}} = (\text{speed})_{\text{Anion}}$$

→ salt bridge ⟶ liquid-liquid junction potential (l.J.P.) Ko Khatam Krta hai

(abhi explain nahi Kiya hai)



Galvanic cell

(a) Daniel cell

(ii) Electrode

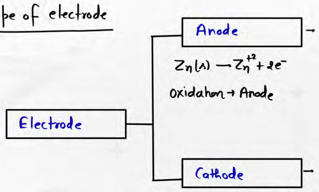
It is a Conductor device, which Connect inner circuit and outer circuit of cell.

$Zn(s) \rightarrow Zn^{+2} + 2e^-$

$Cu^{+2}(aq) + 2e^- \rightarrow Cu(s)$

Ex → Electrode → metal rod, wire, layer

Type of electrode



Electrode

Anode → If electrode oxidise during working of cell

$Zn(s) \rightarrow Zn^{+2} + 2e^-$

Oxidation → Anode

Cathode → If electrode self Reduce

$Cu^{+2}(aq) + 2e^- \rightarrow Cu(s)$

Reduction → Cathode

Q select the correct naming of given halfcell

$\overset{0}{Ni}(s) \rightarrow \overset{+2}{Ni^{+2}}(aq) + 2e^-$

[✓ (a) Anode]

(b) Cathode

Q → $2\overset{-1}{Cl^-} \rightarrow \overset{0}{Cl_2}$

halfcell is

[✓ (a) Anode]

(b) Cathode

Q $Ag^+(aq) + e^- \rightarrow Ag(s)$

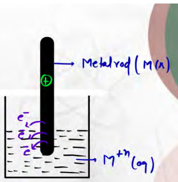
(a) Anode

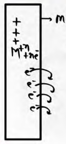
[✓ (b) Cathode]

Type of electrode

(A) Metal- Metal ion electrode (halfcell)

Metal rod (M(s)) + electrolyte ($M^{+n}(aq)$)





metal rod

$M \rightarrow (M^{+n} + ne^-)$

$M(s) \rightarrow M^{+n}(aq) + ne^- \rightarrow$

Ex → $Zn(s) \rightarrow Zn^{+2} + 2e^-$

$Ni(s) \rightarrow Ni^{+2}(aq) + 2e^-$

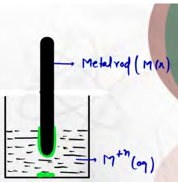
$Ag(s) \rightarrow Ag^+(aq)$

Anode
Oxidation

Type of electrode

(A) Metal- Metal ion electrode (halfcell)

Metal rod (M(s)) + electrolyte ($M^{+n}(aq)$)



metal rod

$\overset{+n}{M^{+n}}(aq) + e^- \rightarrow \overset{0}{M}(s) \rightarrow$ Reduction → Cathode

$Fe^{+2}(aq) + 2e^- \rightarrow Fe(s)$

$Cu^{+2}(aq) + 2e^- \rightarrow Cu(s)$

Cathode

Reduction

Q select Cathode metal-metal ion half cell

(a) ✗ $Zn(s) \longrightarrow \overset{+2}{Zn}(aq) + 2e^-$

(b) ✓ $\overset{+1}{Ag^+}(aq) + e^- \longrightarrow \overset{0}{Ag}(s)$

(c) ✗ $Ni^{+2}(aq) \longrightarrow \overset{+4}{Ni}(aq) + 2e^-$

(d) ✗ $\overset{0}{Cl_2}(g) \longrightarrow \overset{-1}{Cl^-}(aq)$

Red = Cathode

$\overset{+2}{Cu^{+2}} + \overset{0}{Zn} \longrightarrow \overset{+2}{Zn^{+2}} + \overset{0}{Cu}$

Oxd → Anode

select Anode half cell and Cathode half rxn.

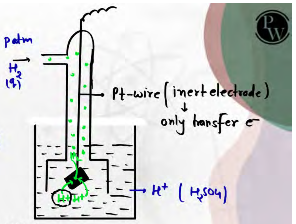
(a) $Cu^{+2} \longrightarrow Cu$, $Cu(s) \longrightarrow Cu^{+2}(aq)$ ✗

(b) ✓ $Zn \longrightarrow Zn^{+2}$, $Cu^{+2} \longrightarrow Cu$

(C) None of these

(B) Gas-Gas ion half cell (Electrode)

Gas	Gas ion (Electrode)
H_2 →	H^+ (HCl, H_2SO_4, HNO_3)
Cl_2 →	Cl^- ($NaCl$, KCl)
O_2 →	OH^- ($NaOH$, KOH)



Anode → $H_{2(g)} \longrightarrow 2H^+_{(aq)} + 2e^-$

Cathode → $2H^+_{(aq)} + 2e^- \longrightarrow H_{2(g)}\uparrow$

Ex → select gas-gas ion halfcell

(a) $Cl_2(g) \longrightarrow 2Cl^-(g)$

(b) $O_2 \longrightarrow OH^-(aq)$

(c) $2H^+(aq) \longrightarrow H_2(g)$

(d) ✓ All

Ex → select correct matching

Ⓐ → $Zn(s) \longrightarrow Zn^{+2}(aq) + 2e^-$

(B) → $H_2(g) \longrightarrow 2H^+(aq) + 2e^-$

(a) A → Metal-metal ion halfcell

(b) B → metal-metal ion halfcell

(c) B → Gas-Gas ion halfcell

(d) ✓ 'A' and 'c' Correct

Ex → $\overset{0}{Fe}(s) + 2\overset{+1}{H^+}(aq) \longrightarrow \overset{+2}{Fe^{+2}}(aq) + \overset{0}{H_2}(g)$

Anode

Red → cathode

select correct statements

(a) ✓ Anode → Metal-Metal ions electrode

(b) ✗ Anode → Gas-Gas ions electrode

(c) ✗ Cathode → metal-metal ion electrode

(d) None of these

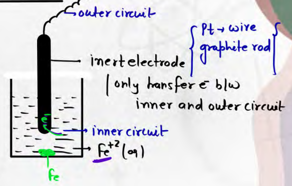
(3) Oxidation - Reduction halfcell (Electrode)

→ Electrolyte-oxidise and Redue

$Fe^{+2}_{(aq)} + 2e^- \longrightarrow Fe(s) \rightarrow$ Cathode

$Fe^{+2}(aq) \longrightarrow Fe^{+3}(aq) + e^- \rightarrow$ Anode

→ d-Block



Q→ $Zn(s) + Cu^{+2}(aq) \longrightarrow Zn^{+2}(aq) + Cu(s)$
is Daniel cell. Daniel cell is a Combination of ____ 2-half cell

(a) Metal - Metal ion ✓

(b) Gas - Gas ion

(c) oxdn - Red - half cell

(d) None of these

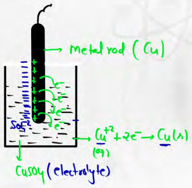
Electrode potential

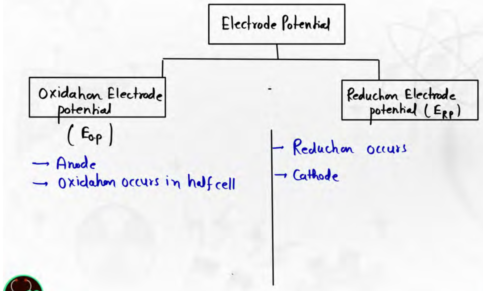
definition →

2→ layer opposite charge

or double layer charge

→ "Potential difference create b/w electrode and electrolyte solution" → Electrode potential





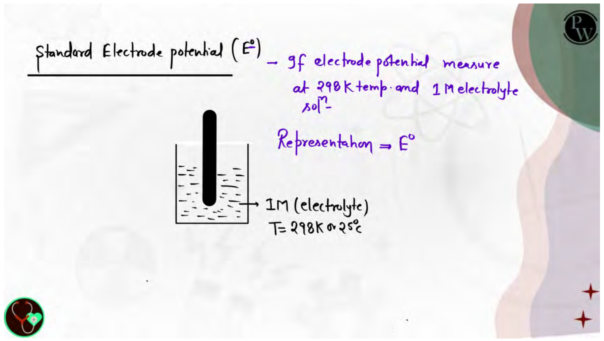
Electrode Potential

Oxidation Electrode potential ($E_{o.p}$)

- → Anode
- → Oxidation occurs in half cell

Reduction Electrode potential (E_{RP})

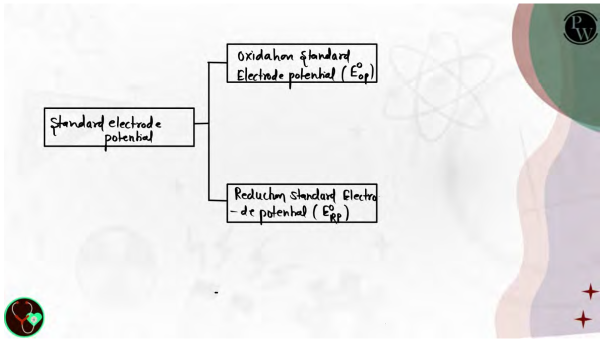
- → Reduction occurs
- → Cathode



Standard Electrode potential (E°)

- If electrode potential measure at 298 K temp. and 1 M electrolyte solⁿ-

Representation ⇒ E°

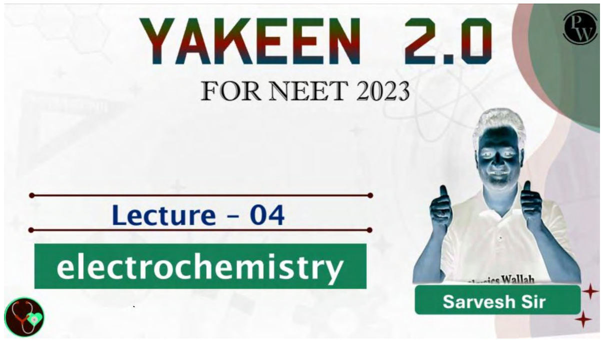


Standard electrode potential

Oxidation standard Electrode potential (E°_{op})

Reduction standard Electro-de potential (E°_{RP})

THANK YOU



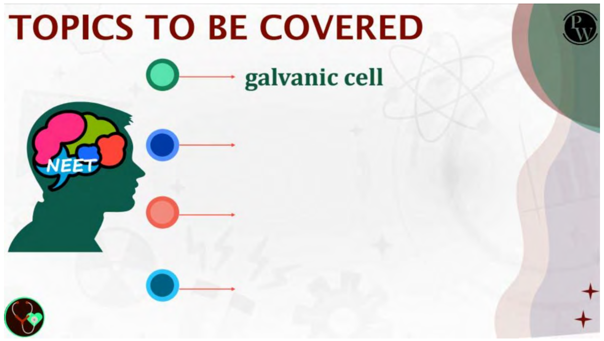
YAKEEN 2.0

FOR NEET 2023

Lecture - 04

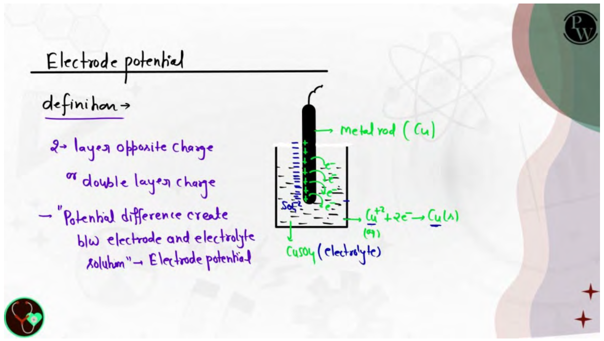
electrochemistry

Sarvesh Sir



TOPICS TO BE COVERED

- galvanic cell



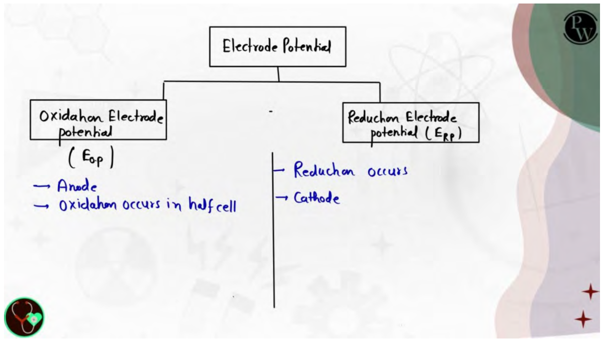
Electrode potential

definition →

2→ layer opposite charge

or double layer charge

→ "Potential difference create b/w electrode and electrolyte solution" → Electrode potential



Electrode Potential

Oxidation Electrode potential ($E_{O.P}$)

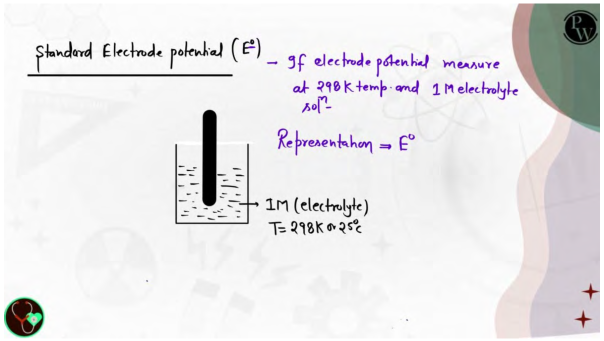
→ Anode

→ Oxidation occurs in half cell

Reduction Electrode potential (E_{RP})

→ Reduction occurs

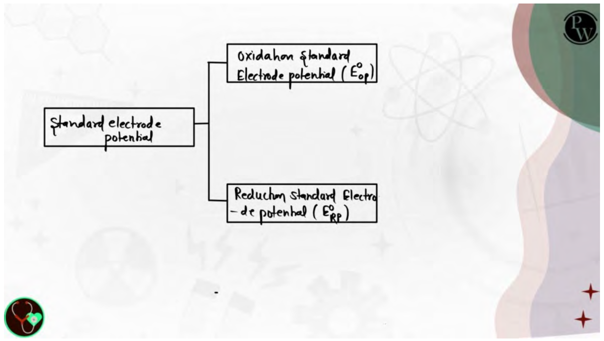
→ Cathode



Standard Electrode potential (E°)

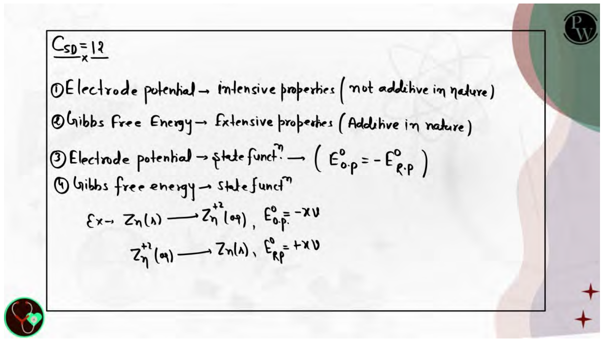
→ If electrode potential measure at 298 K temp. and 1 M electrolyte solⁿ.

Representation ⇒ E°



Standard electrode potential

- Oxidation standard Electrode potential (E^0_{op})
- Reduction standard Electrode potential (E^0_{RP})



$C_{SD} \underset{x}{=} 12$

① Electrode potential → intensive properties (not additive in nature)

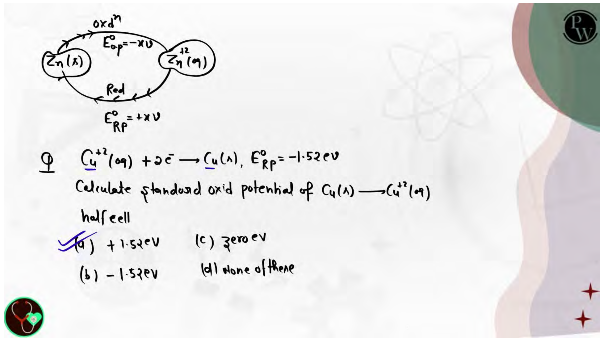
② Gibbs Free Energy → Extensive properties (Additive in nature)

③ Electrode potential → state functⁿ → ($E^0_{O.P} = -E^0_{R.P}$)

④ Gibbs free energy → state functⁿ

Ex → $Zn(s) \longrightarrow Zn^{+2}(aq)$, $E^0_{O.P.} = -xV$

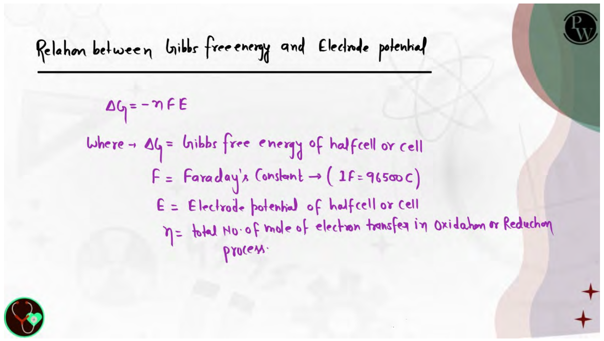
$Zn^{+2}(aq) \longrightarrow Zn(s)$, $E^0_{RP} = +xV$



Q $Cu^{+2}(aq) + 2e^- \longrightarrow Cu(s)$, $E^0_{RP} = -1.52\,eV$

Calculate standard oxid potential of $Cu(s) \longrightarrow Cu^{+2}(aq)$ half cell

[✓ (a)] +1.52 eV
(b) −1.52 eV
(c) zero eV
(d) None of these



Relation between Gibbs free energy and Electrode potential

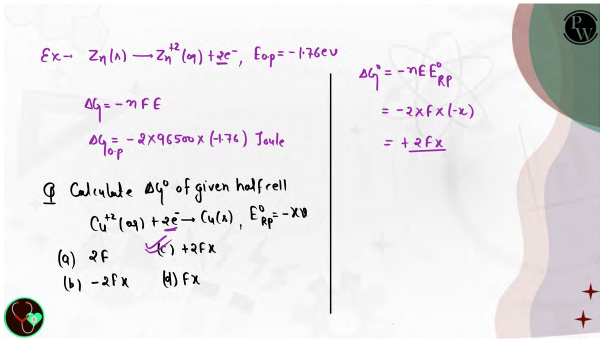
$\Delta G = -nFE$

Where → ΔG = Gibbs free energy of halfcell or cell

F = Faraday's Constant → (1F = 96500 C)

E = Electrode potential of halfcell or cell

n = total No. of mole of electron transfer in Oxidation or Reduction process.



Ex→ $Zn(s) \longrightarrow Zn^{+2}(aq) + \underline{2}e^-$, $E_{op} = -1.76\,eV$

$\Delta G = -nFE$

$\Delta G_{o.p} = -2 \times 96500 \times (-1.76)$ Joule

Q Calculate ΔG° of given halfcell

$Cu^{+2}(aq) + \underline{2}\bar{e} \rightarrow Cu(s)$, $E^\circ_{RP} = -xV$

(a) 2F

[✓ (c) +2Fx]

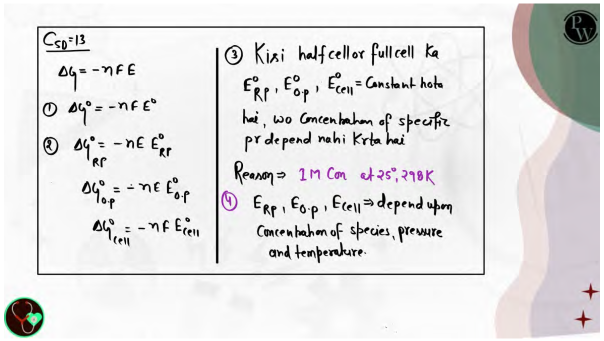
(b) −2Fx

(d) Fx

$\Delta G^\circ = -nE\,E^\circ_{RP}$

$= -2 \times F \times (-x)$

$= \underline{+2Fx}$



$\underline{C_{SD} = 13}$

$\Delta G = -nFE$

① $\Delta G^\circ = -nFE^\circ$

② $\Delta G^\circ_{RP} = -nE\,E^\circ_{RP}$

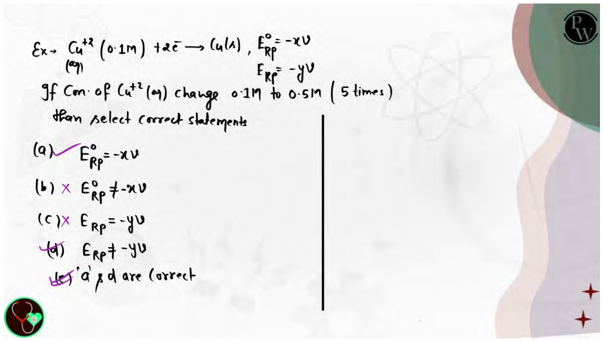
$\Delta G^\circ_{o.p} = -nE\,E^\circ_{o.p}$

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

③ Kisi halfcell or fullcell ka E°_{RP}, $E^\circ_{o.p}$, E°_{cell} = Constant hota hai, wo Concentration of specific pr depend nahi Krta hai

Reason ⇒ 1M Con at 25°, 298K

④ E_{RP}, $E_{o.p}$, E_{cell} ⇒ depend upon Concentration of species, pressure and temperature.



Ex→ $Cu^{+2}_{(aq)}\,(0.1M) + 2\bar{e} \longrightarrow Cu(s)$, $E^\circ_{RP} = -xV$

$E_{RP} = -yV$

If Con. of $Cu^{+2}(aq)$ change 0.1M to 0.5M (5 times) than select correct statements

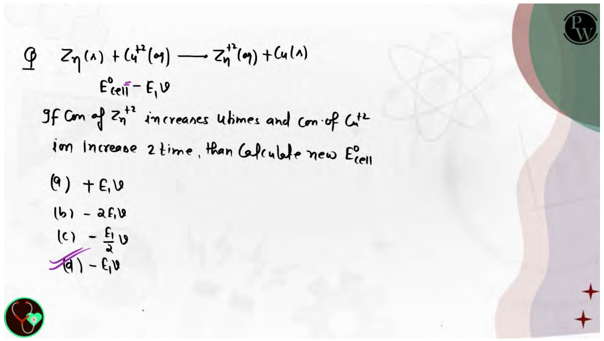
[✓ (a)] $E^\circ_{RP} = -xV$

(b) × $E^\circ_{RP} \neq -xV$

(c) × $E_{RP} = -yV$

[✓ (d)] $E_{RP} \neq -yV$

[✓ (e)] 'a' & d are correct



Q $Zn(s) + Cu^{+2}(aq) \longrightarrow Zn^{+2}(aq) + Cu(s)$

$E^\circ_{cell} = -E_1 V$

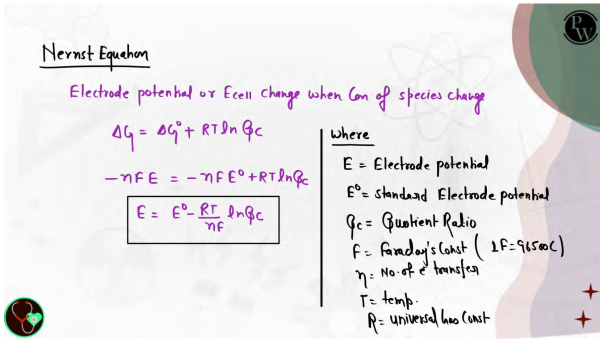
If Con of Zn^{+2} increases 4times and con. of Cu^{+2} ion increase 2 time, than Calculate new E°_{cell}

(a) $+E_1 V$

(b) $-2E_1 V$

(c) $-\frac{E_1}{2} V$

[✓ (d)] $-E_1 V$



Nernst Equation

Electrode potential or Ecell change when Con. of species change

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

$-nFE = -nFE^\circ + RT \ln Q_c$

$$\boxed{E = E^\circ - \frac{RT}{nF} \ln Q_c}$$

Where

E = Electrode potential

E° = Standard Electrode potential

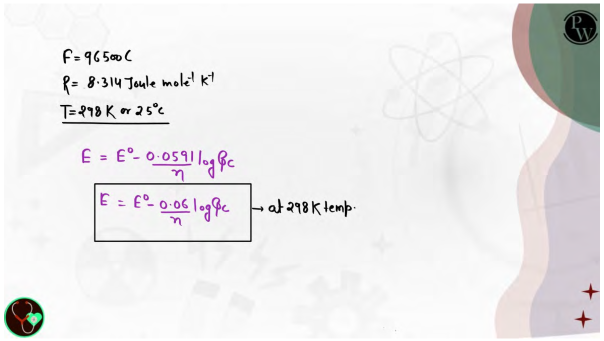
Q_c = Quotient Ratio

F = Faraday's Const ($1F = 96500C$)

n = No. of e^- transfer

T = temp.

R = universal gas Const



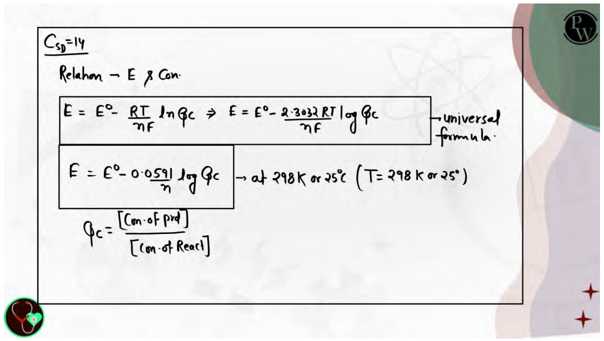
$F = 96500C$

$R = 8.314$ Joule mole^{-1} K^{-1}

$T = 298$ K or 25°c

$E = E^\circ - \frac{0.0591}{n} \log Q_c$

$\boxed{E = E^\circ - \frac{0.06}{n} \log Q_c}$ → at 298 K temp.



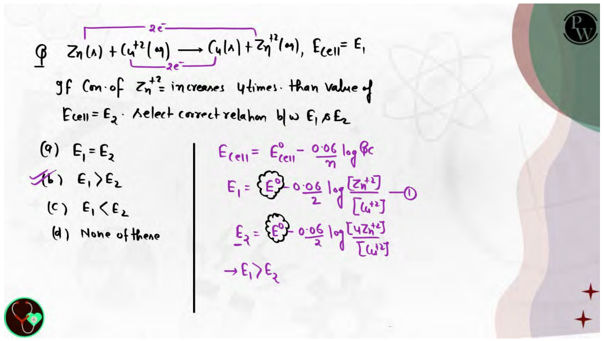
$C_{SD} = 14$

Relation → E & Con.

$\boxed{E = E^\circ - \frac{RT}{nF} \ln Q_c \Rightarrow E = E^\circ - \frac{2.3032\,RT}{nF} \log Q_c}$ → universal formula.

$\boxed{E = E^\circ - 0.\frac{0591}{n} \log Q_c}$ → at 298K or 25°C ($T = 298$ K or 25°)

$Q_c = \frac{[\text{Con. of prd}]}{[\text{Con. of React}]}$



Q $\overbrace{\mathrm{Zn(s)} + \underbrace{\mathrm{Cu^{+2}(aq)} \longrightarrow \mathrm{Cu(s)}}_{2e^-} + \mathrm{Zn^{+2}}}^{2e^-}\mathrm{(aq)}$, $E_{cell} = E_1$

If Con. of Zn^{+2} = increases 4 times. than value of $E_{cell} = E_2$. select correct relation b/w E_1 & E_2

(a) $E_1 = E_2$

[✓ (b) $E_1 > E_2$]

(c) $E_1 < E_2$

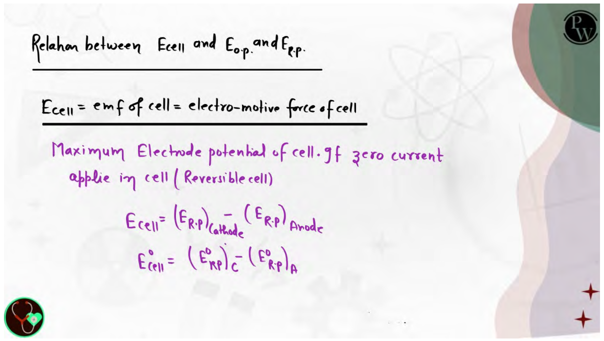
(d) None of these

$E_{cell} = E^\circ_{cell} - \frac{0.06}{n} \log Q_c$

$E_1 = E^\circ - \frac{0.06}{2} \log \frac{[Zn^{+2}]}{[Cu^{+2}]}$ — ①

$E_2 = E^\circ - \frac{0.06}{2} \log \frac{[4Zn^{+2}]}{[Cu^{+2}]}$

→ $E_1 > E_2$



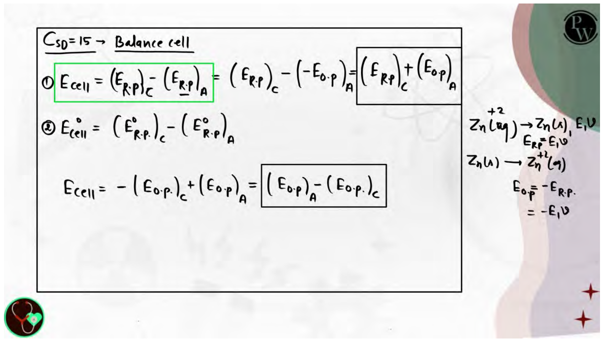
Relation between Ecell and $E_{o.p.}$ and $E_{R.P.}$

E_{cell} = emf of cell = electro-motive force of cell

Maximum Electrode potential of cell. If zero current applie in cell (Reversible cell)

$$E_{cell} = (E_{R.P})_{Cathode} - (E_{R.P})_{Anode}$$

$$E^0_{cell} = (E^0_{R.P})_C - (E^0_{R.P})_A$$



$C_{SD} = 15 \rightarrow$ Balance cell

① $\boxed{E_{cell} = (E_{R.P})_C - (E_{R.P})_A} = (E_{R.P})_C - (-E_{O.P})_A = \boxed{(E_{R.P})_C + (E_{O.P})_A}$

② $E^0_{cell} = (E^0_{R.P.})_C - (E^0_{R.P})_A$

$E_{cell} = -(E_{O.P.})_C + (E_{O.P})_A = \boxed{(E_{O.P})_A - (E_{O.P.})_C}$

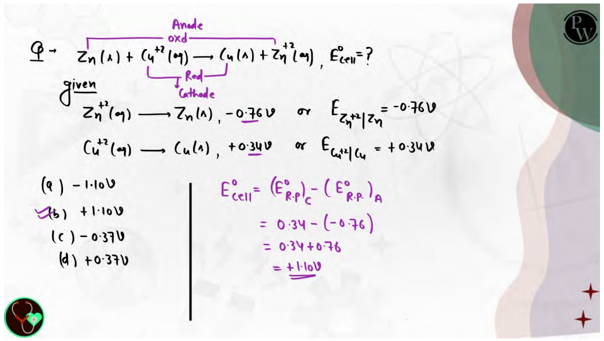
$Zn^{+2}(aq) \rightarrow Zn(s)$, E, V

$E_{RP} = E, V$

$Zn(s) \longrightarrow Zn^{+2}(aq)$

$E_{O.P} = -E_{R.P.}$

$= -E, V$



Anode

oxd

Q → $Zn(s) + Cu^{+2}(aq) \longrightarrow Cu(s) + Zn^{+2}(aq)$, $E^0_{cell} = ?$

Red

Cathode

given

$Zn^{+2}(aq) \longrightarrow Zn(s)$, $-0.76V$ or $E_{Zn^{+2}|Zn} = -0.76V$

$Cu^{+2}(aq) \longrightarrow Cu(s)$, $+0.34V$ or $E_{Cu^{+2}|Cu} = +0.34V$

(a) −1.10V

[✓ (b) +1.10V]

(c) −0.37V

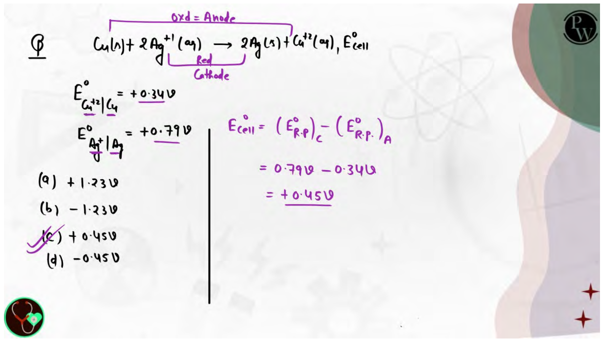
(d) +0.37V

$$E^0_{cell} = (E^0_{R.P})_C - (E^0_{R.P.})_A$$

$$= 0.34 - (-0.76)$$

$$= 0.34 + 0.76$$

$$= \underline{+1.10V}$$



Oxd = Anode

Q $Cu(s) + 2Ag^{+1}(aq) \longrightarrow 2Ag(s) + Cu^{+2}(aq)$, E^0_{cell}

Red

Cathode

$E^0_{Cu^{+2}|Cu} = +0.34V$

$E^0_{Ag^{+}|Ag} = +0.79V$

(a) +1.23V

(b) −1.23V

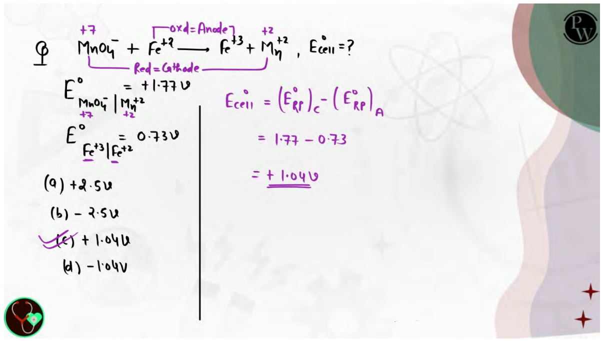
[✓ (c) +0.45V]

(d) −0.45V

$$E^0_{cell} = (E^0_{R.P})_C - (E^0_{R.P.})_A$$

$$= 0.79V - 0.34V$$

$$= \underline{+0.45V}$$



Q $\overset{+7}{MnO_4^-} + Fe^{+2} \longrightarrow Fe^{+3} + \overset{+2}{Mn^{+2}}$, $E^\circ_{cell} = ?$

oxd = Anode

Red = Cathode

$E^\circ_{\overset{}{\underset{+7}{MnO_4^-}} | \underset{+2}{Mn^{+2}}} = +1.77V$

$E^\circ_{Fe^{+3}|Fe^{+2}} = 0.73V$

(a) +2.5V

(b) − 2.5V

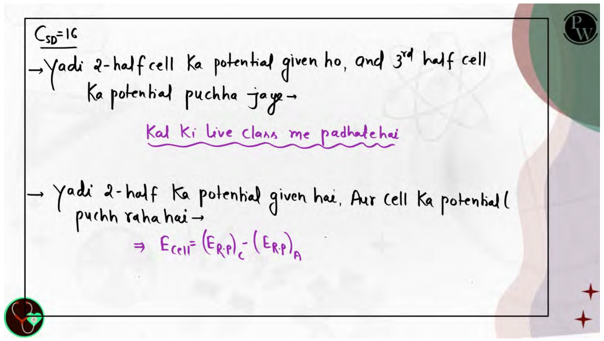
[✓ (c) + 1.04V]

(d) −1.04V

$E^\circ_{cell} = (E^\circ_{RP})_C - (E^\circ_{RP})_A$

$= 1.77 - 0.73$

$= \underline{\underline{+1.04V}}$



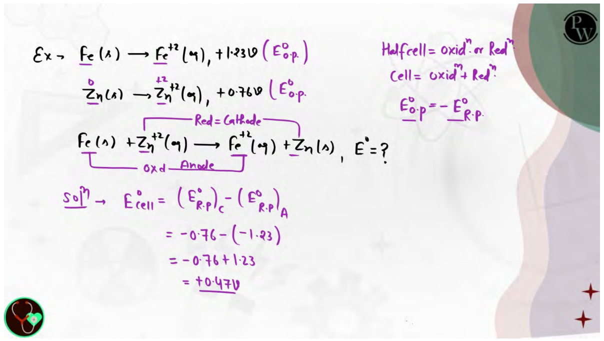
$\underline{C_{SD} = 16}$

→ Yadi 2-half cell Ka potential given ho, and 3rd half cell Ka potential puchha jaye →

Kal Ki Live class me padhatehai

→ Yadi 2-half Ka potential given hai, Aur cell Ka potential puchh raha hai →

$\Rightarrow E_{cell} = (E_{R.P})_C - (E_{R.P})_A$



Ex → $\underline{Fe}(s) \longrightarrow \underline{Fe}^{+2}(aq)$, $+1.23V$ ($E^\circ_{o.p.}$)

$\overset{0}{\underline{Z}}n(s) \longrightarrow \overset{+2}{\underline{Z}}n^{+2}(aq)$, $+0.76V$ ($E^\circ_{o.p.}$

Red = Cathode

$Fe(s) + Zn^{+2}(aq) \longrightarrow Fe^{+2}(aq) + Zn(s)$, $E^\circ = ?$

oxd Anode

Halfcell = oxid^n or Red^n

Cell = oxid^n + Red^n

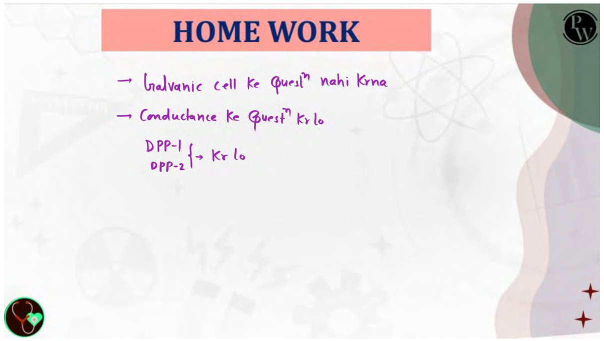
$\underline{E^\circ_{o.p}} = -\underline{E^\circ_{R.P.}}$

$\underline{Sol^n}$ → $E^\circ_{cell} = (E^\circ_{R.P})_C - (E^\circ_{R.P})_A$

$= -0.76 - (-1.23)$

$= -0.76 + 1.23$

$= \underline{+0.47V}$



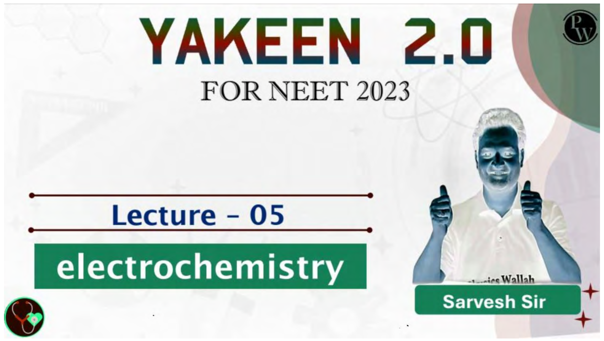
HOME WORK

→ Galvanic cell ke Quest^n nahi Krna

→ Conductance ke Quest^n Kr lo

DPP-1 } → Kr lo
DPP-2 }

THANK YOU



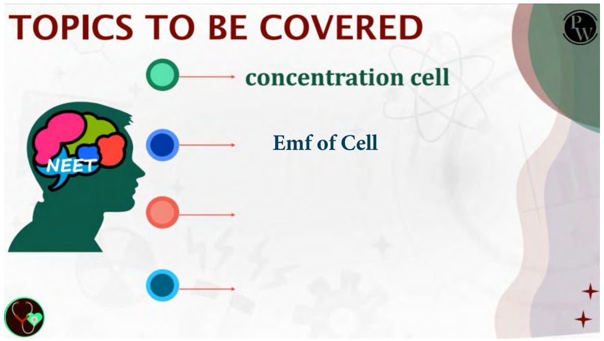
YAKEEN 2.0

FOR NEET 2023

Lecture - 05

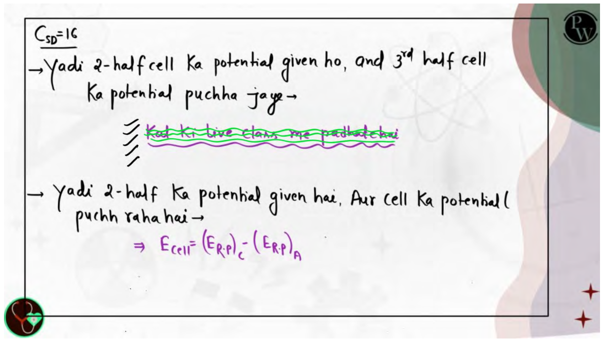
electrochemistry

Sarvesh Sir



TOPICS TO BE COVERED

- concentration cell
- Emf of Cell



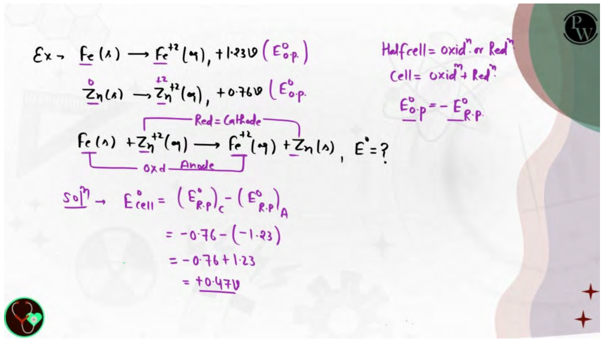
$C_{SD} = 16$

→ Yadi 2-half cell Ka potential given ho, and 3rd half cell Ka potential puchha jaye →

~~Kal Ki live class me padhate hai~~

→ Yadi 2-half Ka potential given hai, Aur cell Ka potential (puchh raha hai →

$\Rightarrow E_{cell} = (E_{R.P})_C - (E_{R.P})_A$



Ex → $\underline{Fe}(s) \longrightarrow \underline{Fe}^{+2}(aq)$, $+1.23V$ ($E^\circ_{o.p.}$)

$\overset{0}{\underline{Zn}}(s) \longrightarrow \overset{+2}{\underline{Zn}}^{+2}(aq)$, $+0.76V$ ($E^\circ_{o.p.}$

Red = Cathode

$\underline{Fe}(s) + \underline{Zn}^{+2}(aq) \longrightarrow \underline{Fe}^{+2}(aq) + \underline{Zn}(s)$, $E^\circ = ?$

oxd Anode

Halfcell = oxid^n or Red^n

Cell = oxid^n + Red^n

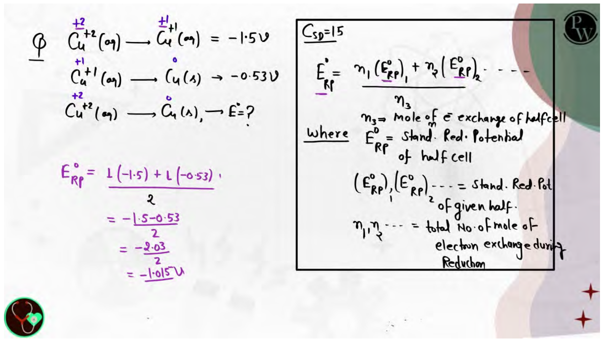
$\underline{E^\circ_{o.p}} = -\underline{E^\circ_{R.P.}}$

$\underline{Sol^n}$ → $E^\circ_{cell} = (E^\circ_{R.P})_C - (E^\circ_{R.P})_A$

$= -0.76 - (-1.23)$

$= -0.76 + 1.23$

$= \underline{+0.47V}$



$\underline{Q}$ $\overset{+2}{Cu}^{+2}(aq) \longrightarrow \overset{+1}{Cu}^{+1}(aq) = -1.5V$

$\overset{+1}{Cu}^{+1}(aq) \longrightarrow \overset{0}{Cu}(s) \rightarrow -0.53V$

$\overset{+2}{Cu}^{+2}(aq) \longrightarrow \overset{0}{Cu}(s), \longrightarrow E^\circ = ?$

$$E^\circ_{RP} = \frac{1(-1.5) + 1(-0.53)}{2}$$

$$= -\frac{1.5 - 0.53}{2}$$

$$= -\frac{2.03}{2}$$

$$= \underline{-1.015V}$$

$\underline{C_{SD}=15}$

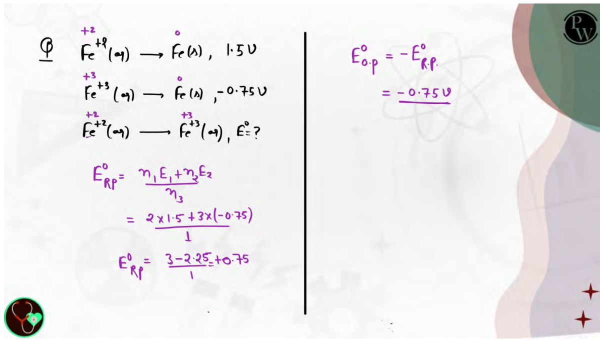
$$E^\circ_{RP} = \frac{n_1(E^\circ_{RP})_1 + n_2(E^\circ_{RP})_2 \cdots}{n_3}$$

n_3 ⇒ mole of $\bar{e}$ exchange of halfcell

$\underline{\text{where}}$ E°_{RP} = Stand. Red. Potential of half cell

$(E^\circ_{RP})_1, (E^\circ_{RP})_2 \cdots$ = Stand. Red. Pot of given half.

$n_1, n_2 \cdots$ = total No. of mole of electron exchange during Reduction



$\underline{Q}$ $\overset{+2}{Fe}^{+2}(aq) \longrightarrow \overset{0}{Fe}(s)$, $1.5V$

$\overset{+3}{Fe}^{+3}(aq) \longrightarrow \overset{0}{Fe}(s)$, $-0.75V$

$\overset{+2}{Fe}^{+2}(aq) \longrightarrow \overset{+3}{Fe}^{+3}(aq)$, $E^\circ = ?$

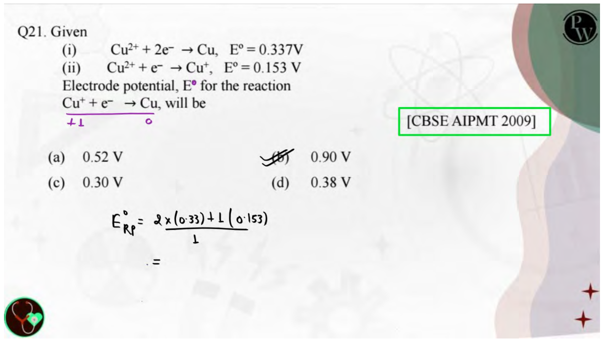
$$E^\circ_{RP} = \frac{n_1E_1 + n_2E_2}{n_3}$$

$$= \frac{2 \times 1.5 + 3 \times (-0.75)}{1}$$

$$E^\circ_{RP} = \frac{3 - 2.25}{1} = +0.75$$

$E^\circ_{o.p} = -E^\circ_{R.P.}$

$= \underline{-0.75V}$



Q21. Given

(i) $Cu^{2+} + 2e^- \rightarrow Cu$, $E^o = 0.337V$

(ii) $Cu^{2+} + e^- \rightarrow Cu^+$, $E^o = 0.153$ V

Electrode potential, E° for the reaction

$\underline{\overset{}{Cu^+} + e^- \rightarrow Cu}$, will be

+1 0

[CBSE AIPMT 2009]

(a) 0.52 V

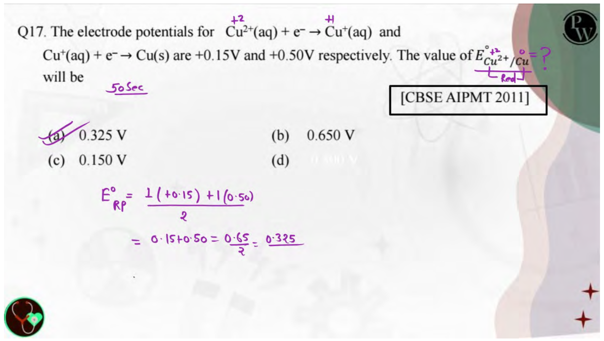
~~(b)~~ 0.90 V

(c) 0.30 V

(d) 0.38 V

$$E^\circ_{RP} = \frac{2 \times (0.33) + 1(0.153)}{1}$$

$=$



Q17. The electrode potentials for $\overset{+2}{Cu^{2+}}(aq) + e^- \rightarrow \overset{+1}{Cu^+}(aq)$ and $Cu^+(aq) + e^- \rightarrow Cu(s)$ are +0.15V and +0.50V respectively. The value of $\underbrace{E^\circ_{\overset{+2}{Cu^{2+}}/\overset{0}{Cu}}}_{Red} = ?$ will be

50 Sec

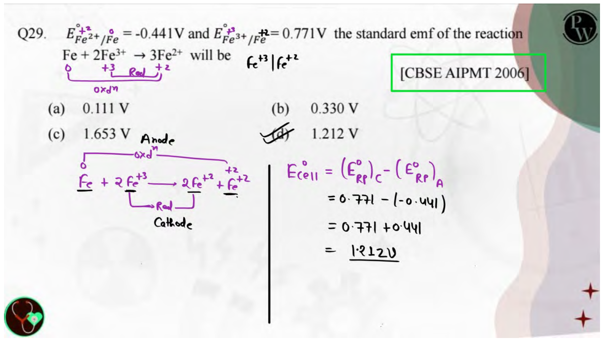
[CBSE AIPMT 2011]

(a) 0.325 V ✓ (b) 0.650 V

(c) 0.150 V (d)

$$E^0_{RP} = \frac{1(+0.15) + 1(0.50)}{2}$$

$$= 0.15 + 0.50 = \frac{0.65}{2} = \underline{0.325}$$



Q29. $E^\circ_{\overset{+2}{Fe^{2+}}/\overset{0}{Fe}} = -0.441V$ and $E^\circ_{\overset{+3}{Fe^{3+}}/\overset{+2}{Fe}} = 0.771V$ the standard emf of the reaction $\underbrace{\overset{0}{Fe} + \underbrace{2\overset{+3}{Fe^{3+}} \rightarrow 3\overset{+2}{Fe^{2+}}}_{Red}}_{oxd^n}$ will be $Fe^{+3}|Fe^{+2}$

[CBSE AIPMT 2006]

(a) 0.111 V (b) 0.330 V

(c) 1.653 V (d) 1.212 V ✓

Anode

oxd^n

$$\overset{0}{Fe} + 2Fe^{+3} \longrightarrow 2Fe^{+2} + \overset{+2}{Fe^{+2}}$$

Red

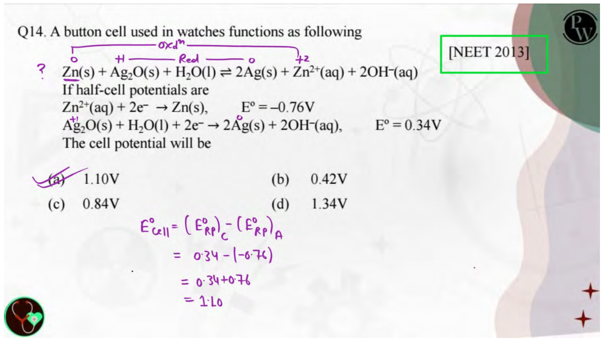
Cathode

$$E^0_{cell} = (E^0_{RP})_C - (E^0_{RP})_A$$

$$= 0.771 - (-0.441)$$

$$= 0.771 + 0.441$$

$$= \underline{1.212V}$$



Q14. A button cell used in watches functions as following

[NEET 2013]

? $\underbrace{\overset{0}{Zn}(s) + \overset{+1}{Ag_2}O(s) + H_2O(l) \rightleftharpoons 2\overset{0}{Ag}(s) + \overset{+2}{Zn^{2+}}}_{oxd^n}(aq) + 2OH^-(aq)$ (Red: $\overset{+1}{Ag} \rightarrow \overset{0}{Ag}$)

If half-cell potentials are

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$, E° = –0.76V

$\overset{+1}{Ag_2}O(s) + H_2O(l) + 2e^- \rightarrow 2\overset{0}{Ag}(s) + 2OH^-(aq)$, E° = 0.34V

The cell potential will be

(a) 1.10V ✓ (b) 0.42V

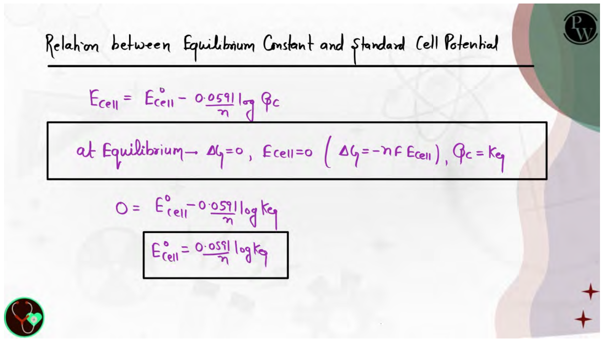
(c) 0.84V (d) 1.34V

$$E^0_{cell} = (E^0_{RP})_C - (E^0_{RP})_A$$

$$= 0.34 - (-0.76)$$

$$= 0.34 + 0.76$$

$$= 1.10$$



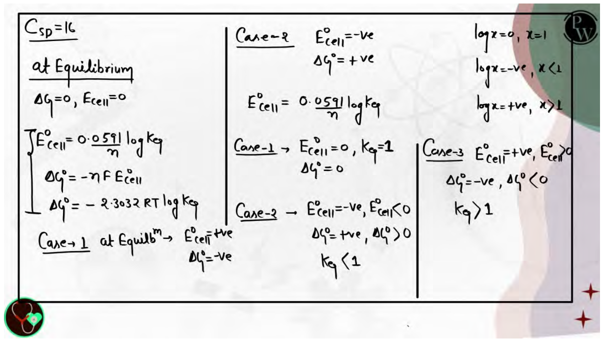
Relation between Equilibrium Constant and Standard Cell Potential

$$E_{cell} = E^0_{cell} - \frac{0.0591}{n} \log Q_C$$

$$\boxed{\text{at Equilibrium} \rightarrow \Delta G = 0,\ E_{cell} = 0\ (\Delta G = -nFE_{cell}),\ Q_C = K_{eq}}$$

$$0 = E^0_{cell} - \frac{0.0591}{n} \log K_{eq}$$

$$\boxed{E^0_{cell} = \frac{0.0591}{n} \log K_{eq}}$$



$C_{sp}=16$

at Equilibrium

$\Delta G=0,\ E_{cell}=0$

$E^\circ_{cell} = 0.0591\,\frac{}{n}\log K_{eq}$ → $E^\circ_{cell} = \frac{0.0591}{n}\log K_{eq}$

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

$\Delta G^\circ = -2.3032\,RT\log K_{eq}$

Case→1 at Equilb^m → $E^\circ_{cell} = +ve$, $\Delta G^\circ = -ve$

Case-2 $E^\circ_{cell} = -ve$, $\Delta G^\circ = +ve$

$E^\circ_{cell} = \frac{0.0591}{n}\log K_{eq}$

Case-1 → $E^\circ_{cell} = 0,\ K_{eq} = 1$, $\Delta G^\circ = 0$

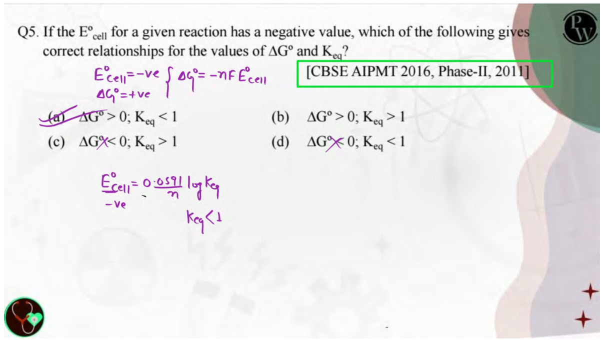
Case-2 → $E^\circ_{cell} = -ve,\ E^\circ_{cell} < 0$, $\Delta G^\circ = +ve,\ \Delta G^\circ > 0$, $K_{eq} < 1$

$\log x = 0,\ x = 1$

$\log x = -ve,\ x < 1$

$\log x = +ve,\ x > 1$

Case-3 $E^\circ_{cell} = +ve,\ E^\circ_{cell} > 0$, $\Delta G^\circ = -ve,\ \Delta G^\circ < 0$, $K_{eq} > 1$



Q5. If the E°_{cell} for a given reaction has a negative value, which of the following gives correct relationships for the values of ΔG° and K_{eq}?

[CBSE AIPMT 2016, Phase-II, 2011]

$E^\circ_{cell} = -ve$, $\Delta G^\circ = +ve$ $\Big\{$ $\Delta G^\circ = -nFE^\circ_{cell}$

(a) ✓ $\Delta G^\circ > 0;\ K_{eq} < 1$

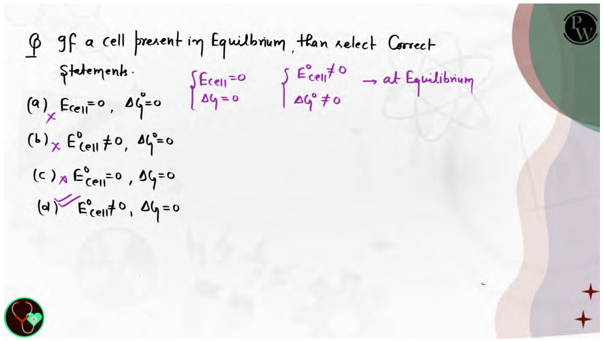
(b) $\Delta G^\circ > 0;\ K_{eq} > 1$

(c) ΔG° ✗ $< 0;\ K_{eq} > 1$

(d) ΔG° ✗ $< 0;\ K_{eq} < 1$

$\underset{-ve}{E^\circ_{cell}} = \frac{0.0591}{n}\log K_{eq}$

$K_{eq} < 1$



Q If a cell present in Equilibrium, than select Correct statements.

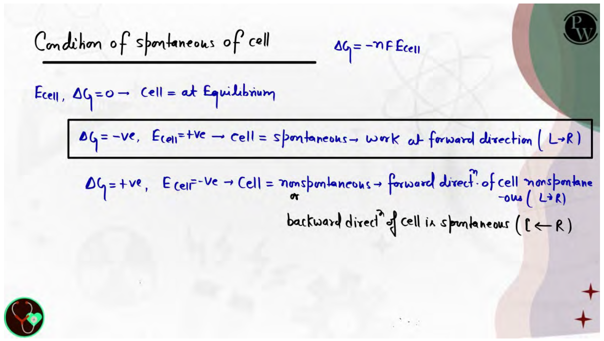
(a) ✗ $E_{cell} = 0,\ \Delta G^\circ = 0$

(b) ✗ $E^\circ_{cell} \neq 0,\ \Delta G^\circ = 0$

(c) ✗ $E^\circ_{cell} = 0,\ \Delta G = 0$

(d) ✓ $E^\circ_{cell} \neq 0,\ \Delta G = 0$

$\Big\{$ $E_{cell} = 0$, $\Delta G = 0$ $\quad\Big\{$ $E^\circ_{cell} \neq 0$, $\Delta G^\circ \neq 0$ → at Equilibrium



Condition of spontaneous of cell

$\Delta G = -nFE_{cell}$

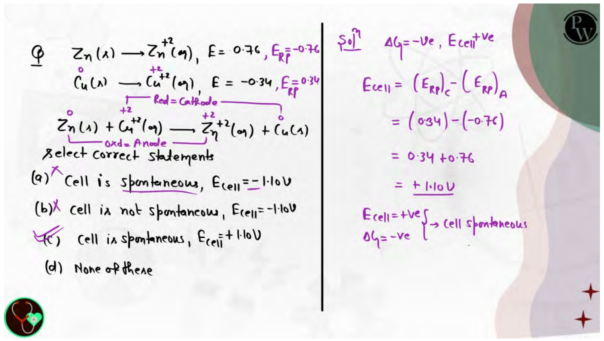
$E_{cell},\ \Delta G = 0$ → Cell = at Equilibrium

$\boxed{\Delta G = -ve,\ E_{cell} = +ve \rightarrow \text{cell = spontaneous} \rightarrow \text{work at forward direction } (L \rightarrow R)}$

$\Delta G = +ve$, $E_{cell} = -ve$ → Cell = nonspontaneous → forward direct^n. of cell nonspontaneous (L→R)

or

backward direct^n of cell is spontaneous (L ← R)



Q $Zn(s) \longrightarrow Zn^{+2}(aq)$, $E = 0.76$, $E_{RP} = -0.76$

$\overset{0}{Cu}(s) \longrightarrow \overset{+2}{Cu^{+2}}(aq)$, $E = -0.34$, $E_{RP} = 0.34$

Red = Cathode

$\overset{0}{Zn}(s) + \overset{+2}{Cu^{+2}}(aq) \longrightarrow \overset{+2}{Zn^{+2}}(aq) + \overset{0}{Cu}(s)$

oxd = Anode

Select correct statements

(a) ✗ Cell is spontaneous, $E_{cell} = -1.10V$

(b) ✗ cell is not spontaneous, $E_{cell} = -1.10V$

✓ (c) cell is spontaneous, $E_{cell} = +1.10V$

(d) None of these

Solⁿ $\Delta G = -ve$, $E_{cell} +ve$

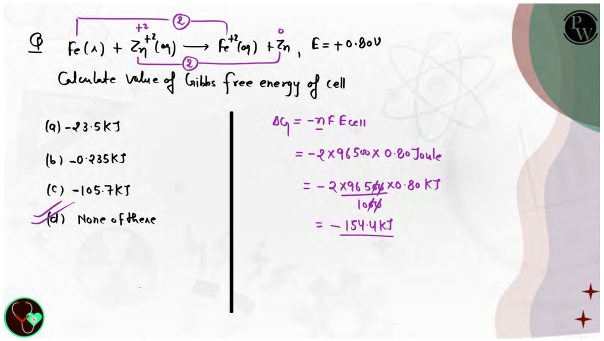
$E_{cell} = (E_{RP})_C - (E_{RP})_A$

$= (0.34) - (-0.76)$

$= 0.34 + 0.76$

$= \underline{+1.10V}$

$\left.\begin{array}{l} E_{cell} = +ve \\ \Delta G = -ve \end{array}\right\} \to$ cell spontaneous



Q $Fe(s) + \overset{+2}{Zn^{+2}}(aq) \longrightarrow Fe^{+2}(aq) + \overset{0}{Zn}$, $E = +0.80V$

② ②

Calculate value of Gibbs free energy of cell

(a) −23.5KJ

(b) −0.235KJ

(c) −105.7KJ

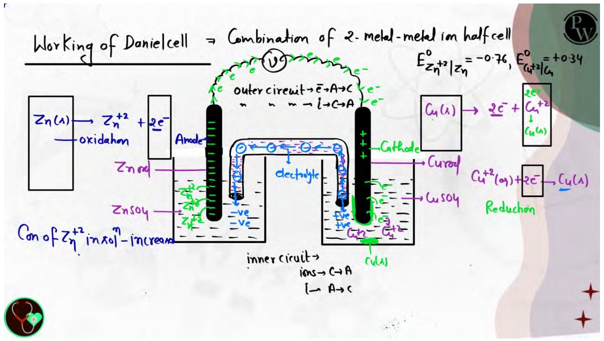
✓ (d) None of these

$\Delta G = -nFE_{cell}$

$= -2 \times 96500 \times 0.80$ Joule

$= -\frac{2 \times 965\cancel{00}}{10\cancel{00}} \times 0.80$ KJ

$= \underline{-154.4KJ}$



Working of Danielcell ⇒ Combination of 2-metal-metal ion halfcell

$E^0_{Zn^{+2}|Zn} = -0.76$, $E^0_{Cu^{+2}|Cu} = +0.34$

$Zn(s) \longrightarrow Zn^{+2} + \underline{2e^-}$

oxidation

$Cu(s) \longrightarrow \underline{2e^-} + Cu^{+2}$

$Cu^{+2}(aq) + 2e^- \longrightarrow Cu(s)$

Reduction

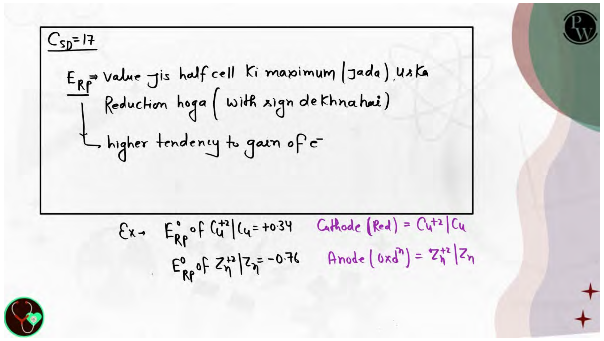
outer circuit → $\bar{e}$ → A → C

" " " → I → C → A

inner circuit → ions → C → A

I → A → C

Con of Zn^{+2} in solⁿ − increase



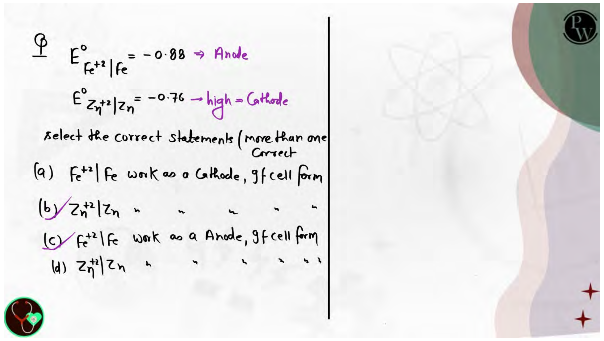
$C_{SD} = 17$

$\underline{E_{RP}}$ ⇒ value jis half cell ki maximum (jada), uska Reduction hoga (with sign dekhna hai)

↳ higher tendency to gain of e⁻

Ex → E^0_{RP} of $Cu^{+2}|Cu = +0.34$ Cathode (Red) = $Cu^{+2}|Cu$

E^0_{RP} of $Zn^{+2}|Zn = -0.76$ Anode (oxdⁿ) = $Zn^{+2}|Zn$



Q $E^{0}_{Fe^{+2}|Fe} = -0.88$ ⇒ Anode

$E^{0}_{Zn^{+2}|Zn} = -0.76$ → high ⇒ Cathode

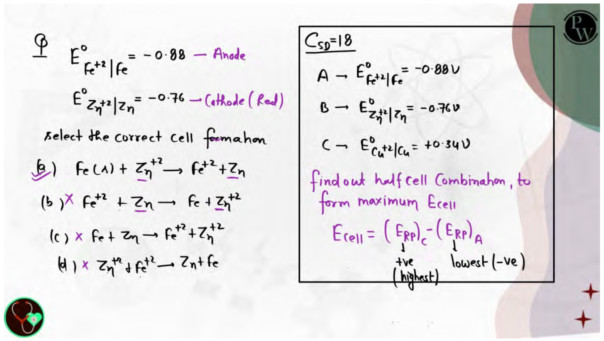
select the correct statements (more than one Correct

(a) $Fe^{+2}|Fe$ work as a Cathode, If cell form

(b) ✓ $Zn^{+2}|Zn$ " " " " "

(c) ✓ $Fe^{+2}|Fe$ work as a Anode, If cell form

(d) $Zn^{+2}|Zn$ " " " " " "



Q $E^{0}_{Fe^{+2}|Fe} = -0.88$ — Anode

$E^{0}_{Zn^{+2}|Zn} = -0.76$ → Cathode (Red)

select the correct cell formation

(a) ✓ $Fe(s) + \underline{Zn}^{+2} \longrightarrow Fe^{+2} + \underline{Zn}$

(b) ✗ $Fe^{+2} + \underline{Zn} \longrightarrow Fe + \underline{Zn}^{+2}$

(c) ✗ $Fe + Zn \longrightarrow Fe^{+2} + Zn^{+2}$

(d) ✗ $Zn^{+2} + Fe^{+2} \longrightarrow Zn + Fe$

$\underline{C_{SD}=18}$

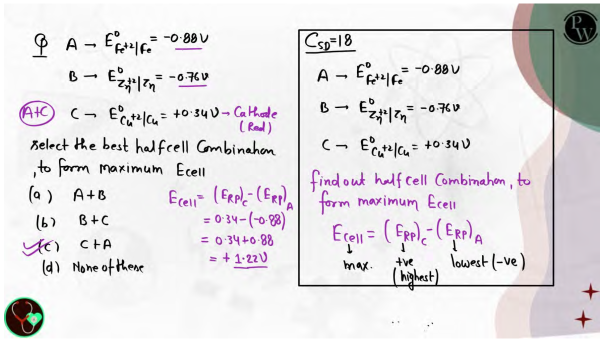
A → $E^{0}_{Fe^{+2}|Fe} = -0.88V$

B → $E^{0}_{Zn^{+2}|Zn} = -0.76V$

C → $E^{0}_{Cu^{+2}|Cu} = +0.34V$

find out half cell Combination, to form maximum Ecell

$$E_{cell} = \underset{\substack{\downarrow \\ +ve \\ (highest)}}{(E_{RP})_C} - \underset{\substack{\downarrow \\ lowest\ (-ve)}}{(E_{RP})_A}$$



Q A → $E^{0}_{Fe^{+2}|Fe} = \underline{-0.88V}$

B → $E^{0}_{Zn^{+2}|Zn} = \underline{-0.76V}$

(A+C) C → $E^{0}_{Cu^{+2}|Cu} = +0.34V$ → Cathode (Red)

select the best halfcell Combination, to form maximum Ecell

(a) A+B

(b) B+C

✓(c) C+A

(d) None of these

$$E_{cell} = (E_{RP})_C - (E_{RP})_A$$
$$= 0.34 - (-0.88)$$
$$= 0.34 + 0.88$$
$$= +\underline{1.22V}$$

$\underline{C_{SD}=18}$

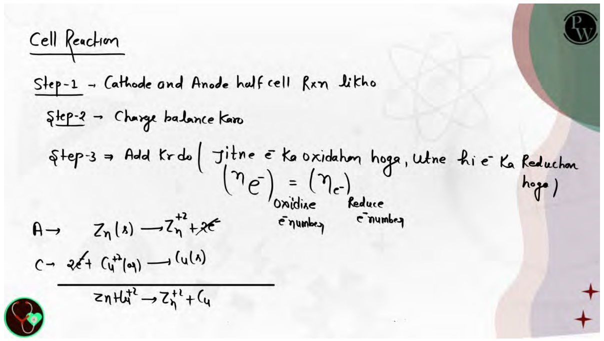
A → $E^{0}_{Fe^{+2}|Fe} = -0.88V$

B → $E^{0}_{Zn^{+2}|Zn} = -0.76V$

C → $E^{0}_{Cu^{+2}|Cu} = +0.34V$

find out half cell Combination, to form maximum Ecell

$$\underset{\substack{\downarrow \\ max.}}{E_{cell}} = \underset{\substack{\downarrow \\ +ve \\ (highest)}}{(E_{RP})_C} - \underset{\substack{\downarrow \\ lowest\ (-ve)}}{(E_{RP})_A}$$



Cell Reaction

Step-1 → Cathode and Anode half cell Rxn likho

Step-2 → Charge balance Karo

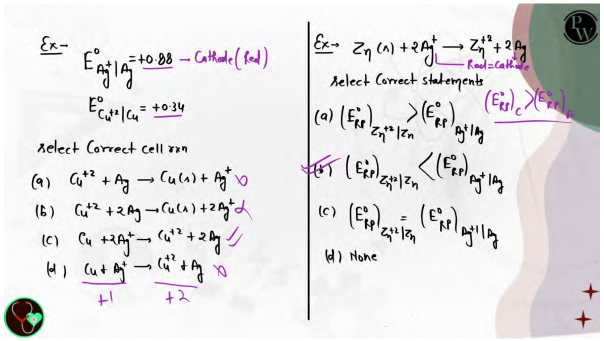
Step-3 ⇒ Add Kr do (Jitne e^- Ka oxidation hoga, utne hi e^- Ka Reduction hoga)

$$\underset{\substack{Oxidise \\ e^-\ number}}{(n_{e^-})} = \underset{\substack{Reduce \\ e^-\ number}}{(n_{e^-})}$$

A → $Zn(s) \longrightarrow Zn^{+2} + \cancel{2e^-}$

C → $\cancel{2e^-} + Cu^{+2}(aq) \longrightarrow Cu(s)$

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



Ex→

$E^0_{Ag^+|Ag} = \underline{+0.88}$ → Cathode (Red)

$E^0_{Cu^{+2}|Cu} = \underline{+0.34}$

select Correct cell rxn

(a) $Cu^{+2} + Ag \rightarrow Cu(s) + Ag^+$ ✗

(b) $Cu^{+2} + 2Ag \rightarrow Cu(s) + 2Ag^+$ ✗

(c) $Cu + 2Ag^+ \rightarrow Cu^{+2} + 2Ag$ ✓

(d) $\underbrace{Cu + Ag^+}_{+1} \rightarrow \underbrace{Cu^{+2} + Ag}_{+2}$ ✗

Ex→ $Zn(s) + 2Ag^+ \rightarrow Zn^{+2} + 2Ag$ Red = Cathode

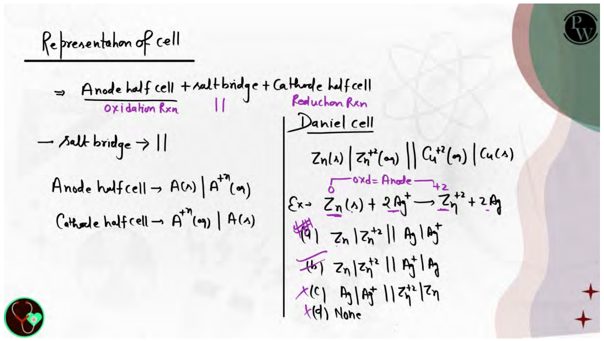
select Correct statements

(a) $(E^0_{RP})_{Zn^{+2}|Zn} > (E^0_{RP})_{Ag^+|Ag}$ $\underline{(E^0_{RP})_C > (E^0_{RP})_A}$

✓(b) $(E^0_{RP})_{Zn^{+2}|Zn} < (E^0_{RP})_{Ag^+|Ag}$

(c) $(E^0_{RP})_{Zn^{+2}|Zn} = (E^0_{RP})_{Ag^{+1}|Ag}$

(d) None



Representation of cell

⇒ $\underbrace{\text{Anode half cell}}_{\text{Oxidation Rxn}}$ + $\underbrace{\text{salt bridge}}_{||}$ + $\underbrace{\text{Cathode half cell}}_{\text{Reduction Rxn}}$

→ salt bridge → ||

Anode half cell → $A(s) \,|\, A^{+n}(aq)$

Cathode half cell → $A^{+n}(aq) \,|\, A(s)$

Daniel cell

$Zn(s) \,|\, Zn^{+2}(aq) \,||\, Cu^{+2}(aq) \,|\, Cu(s)$

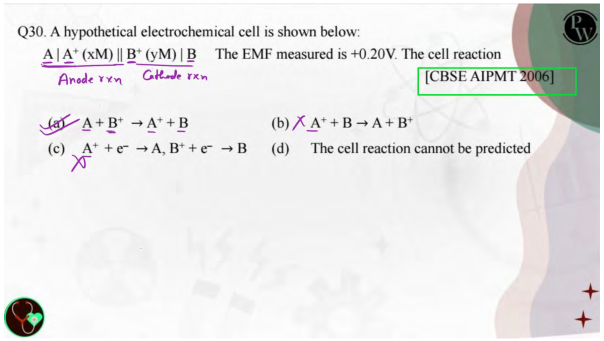
Ex→ $\overset{0}{Zn}(s) + 2Ag^+ \rightarrow \overset{+2}{Zn^{+2}} + 2Ag$ oxd = Anode

(a) $Zn \,|\, Zn^{+2} \,||\, Ag \,|\, Ag^+$

✓(b) $Zn \,|\, Zn^{+2} \,||\, Ag^+ \,|\, Ag$

✗(c) $Ag \,|\, Ag^+ \,||\, Zn^{+2} \,|\, Zn$

✗(d) None

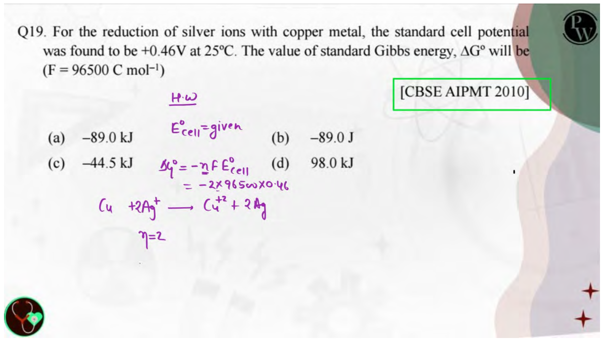


Q30. A hypothetical electrochemical cell is shown below:

$\underbrace{A \,|\, A^+ (xM)}_{\text{Anode rxn}} \,||\, \underbrace{B^+ (yM) \,|\, B}_{\text{Cathode rxn}}$ The EMF measured is +0.20V. The cell reaction

[CBSE AIPMT 2006]

✓(a) $A + B^+ \rightarrow A^+ + B$
(b) ✗ $A^+ + B \rightarrow A + B^+$
(c) $A^+ + e^- \rightarrow A,\ B^+ + e^- \rightarrow B$ ✗
(d) The cell reaction cannot be predicted



Q19. For the reduction of silver ions with copper metal, the standard cell potential was found to be +0.46V at 25ºC. The value of standard Gibbs energy, ΔGº will be (F = 96500 C mol⁻¹)

[CBSE AIPMT 2010]

H.W

E^0_{cell} = given

(a) –89.0 kJ
(b) –89.0 J
(c) –44.5 kJ
(d) 98.0 kJ

$\Delta G^0 = -nFE^0_{cell}$
$= -2 \times 96500 \times 0.46$

$Cu + 2Ag^+ \longrightarrow Cu^{+2} + 2Ag$

$n = 2$

Q1. In the electrochemical cell Zn || $ZnSO_4$ (0.01M) || $CuSO_4$(1.0M) Cu, the emf of this Daniel cell is E_1. When the concentration $ZnSO_4$ is changed to 1.0M and that of $CuSO_4$ changed to 0.01M, the emf changes to E_2. From the followings, which one is the relationship between E_1 and E_2? (Given, RT/F = 0.059)

[NEET 2017, 2003]

H.W.

(a) $E_1 = E_2$
(b) $E_1 < E_2$
(c) $E_1 > E_2$
(d) $E_2 = 0 \neq E_1$

HOME WORK

→ DPP Krni hai
→ ΔG°
→ E°_{cell}
→

THANK YOU

YAKEEN 2.0

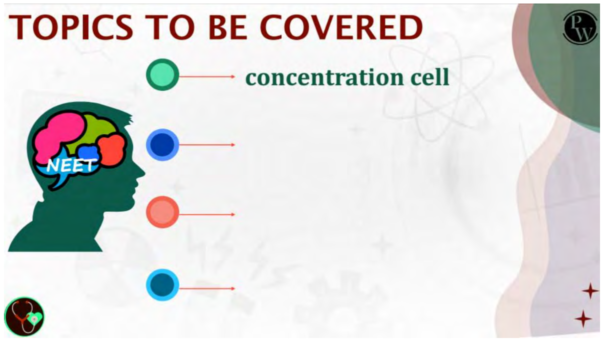
FOR NEET 2023

Qsd →

Lecture – 06

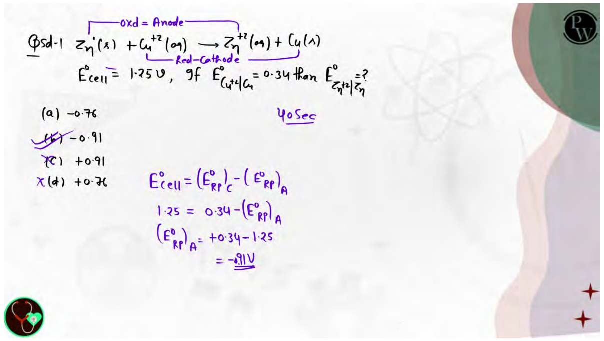
electrochemistry

Sarvesh Sir



TOPICS TO BE COVERED

- concentration cell



$$\text{Qsd-1}\quad \underbrace{Zn(s)}_{} + Cu^{+2}(aq) \longrightarrow Zn^{+2}(aq) + Cu(s)$$

Oxd = Anode

Red-Cathode

$E^\circ_{cell} = 1.25\,V$, If $E^\circ_{Cu^{+2}|Cu} = 0.34$ than $E^\circ_{Zn^{+2}|Zn} = ?$

(a) −0.76

[✓ (b) −0.91]

~~(c)~~ +0.91

x (d) +0.76

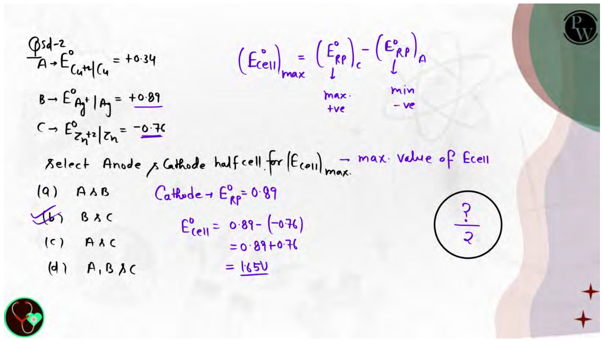
40 Sec

$$E^\circ_{cell} = (E^\circ_{RP})_C - (E^\circ_{RP})_A$$

$$1.25 = 0.34 - (E^\circ_{RP})_A$$

$$(E^\circ_{RP})_A = +0.34 - 1.25$$

$$= \underline{-0.91\,V}$$



Qsd-2

$A \rightarrow E^\circ_{Cu^{+2}|Cu} = +0.34$

$B \rightarrow E^\circ_{Ag^+|Ag} = \underline{+0.89}$

$C \rightarrow E^\circ_{Zn^{+2}|Zn} = \underline{-0.76}$

$$(E^\circ_{cell})_{max} = \underset{\substack{\downarrow \\ \text{max.} \\ +ve}}{(E^\circ_{RP})_C} - \underset{\substack{\downarrow \\ \text{min} \\ -ve}}{(E^\circ_{RP})_A}$$

Select Anode & Cathode half cell for $(E_{cell})_{max.}$ → max. value of Ecell

(a) A & B

[✓ (b) B & C]

(c) A & C

(d) A, B & C

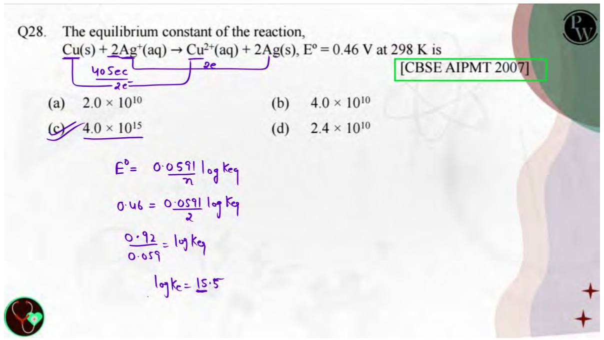
Cathode → $E^\circ_{RP} = 0.89$

$$E^\circ_{cell} = 0.89 - (-0.76)$$

$$= 0.89 + 0.76$$

$$= \underline{1.65\,V}$$

$$\boxed{\frac{?}{2}}$$



Q28. The equilibrium constant of the reaction,
$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$, E° = 0.46 V at 298 K is

2e

40 Sec

2e⁻

[CBSE AIPMT 2007]

(a)	2.0×10^{10}	(b)	4.0×10^{10}
[✓ (c)]	$\underline{4.0 \times 10^{15}}$	(d)	2.4×10^{10}

$$E^\circ = \frac{0.0591}{n} \log K_{eq}$$

$$0.46 = \frac{0.0591}{2} \log K_{eq}$$

$$\frac{0.92}{0.059} = \log K_{eq}$$

$$\log K_c = \underline{15}.5$$

Q37. Cell reaction is spontaneous when

30 Sec

[CBSE AIPMT 2000]

(a) ✗ E_{red} is negative ✗ (b) E_{red} is positive $E_{cell} \rightarrow$ ✗

[✓ (c)] ΔG is negative ✗ (d) ΔG is positive

$\Delta G = -ve \rightarrow$ spontaneous

$\underline{E_{cell}} = +ve$

Q32. The standard EMF of a galvanic cell involving cell reaction with n = 2 is found to be 0.295V at $\underline{25}$ºC. The equilibrium constant of the reaction would be (Given F = $\underline{96500}$C mol^{-1}, $\underline{R = 8.}$314 JK^{-1} mol^{-1})

[CBSE AIPMT 2004]

(a) 2.0×10^{11} (b) 4.0×10^{12}

(c) 1.0×10^{2} [✓ (d)] 1.0×10^{10}

$E^\circ = 0.\frac{0591}{n} \log K_{eq}$

$0.295 = 0.\frac{0591}{2} \log K_{eq}$

$0.295 \times 2 = 0.0591 \log K_{eq}$

$\frac{\cancel{0.59}}{10} = \cancel{0.0591} \log K_{eq}$

$\log K_{eq} = 10$

$\underline{K_{e} = 10^{10}}$

Concentration cell

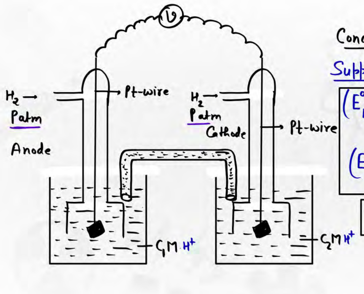
Cell = Anode half + Cathode half cell

Ex → Daniel cell → $Zn + Cu^{+2} \longrightarrow Zn^{+2} + Cu$

$\underbrace{Zn | Zn^{+2}}_{\text{Anode}} || \underbrace{Cu^{+2} | Cu}_{\text{Cathode}}$

Anode half cell
Cathode half cell } → same half cell → Concentration cell

Ex → $Zn | Zn^{+2}(C_1M) || Zn^{+2}(C_2) | Zn(s)$



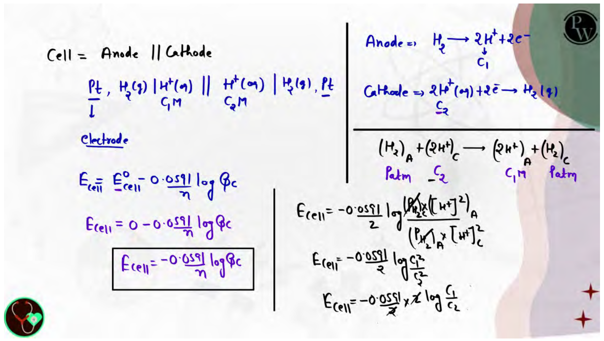
Concentration cell

Suppose

$\left(E^\circ_{RP}\right)_{H^+_{aq}|H_2(g)} = 0$

$\left(E^\circ_{O\cdot P}\right)_{H_2|H^+(aq)} = 0$

$\boxed{E^\circ_{O\cdot P} = -E^\circ_{R\cdot P}}$



Cell = Anode || Cathode

$\underset{\downarrow}{\underline{Pt}},\ H_2(g)\,|\,\underset{C_1M}{H^+(aq)}\ ||\ \underset{C_2M}{H^+(aq)}\,|\,H_2(g),\ \underline{Pt}$

electrode

$E_{cell} = \underline{E}^\circ_{cell} - \frac{0.0591}{n}\log Q_c$

$E_{cell} = 0 - \frac{0.0591}{n}\log Q_c$

$\boxed{E_{cell} = -\frac{0.0591}{n}\log Q_c}$

Anode ⇒ $H_2 \longrightarrow \underset{C_1}{\overset{}{2H^+}} + 2e^-$

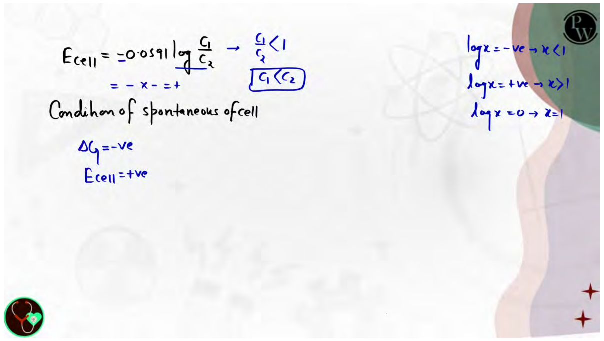
Cathode ⇒ $\underset{C_2}{2H^+}(aq) + 2\bar{e} \longrightarrow H_2(g)$

$\underset{P_{atm}}{(H_2)_A} + \underset{C_2}{(2H^+)_C} \longrightarrow \underset{C_1M}{(2H^+)_A} + \underset{P_{atm}}{(H_2)_C}$

$E_{cell} = -\frac{0.0591}{2}\log\frac{\cancel{(P_{H_2})_C}\times([H^+]^2)_A}{\cancel{(P_{H_2})_A}\times[H^+]^2_C}$

$E_{cell} = -\frac{0.0591}{2}\log\frac{C_1^2}{C_2^2}$

$E_{cell} = -\frac{0.0591}{\cancel{2}}\times\cancel{2}\log\frac{C_1}{C_2}$



$E_{cell} = \underline{-}0.0591\log\frac{C_1}{C_2} \rightarrow \frac{C_1}{C_2} < 1$

$= - \times - = +$

$\boxed{C_1 < C_2}$

Condition of spontaneous of cell

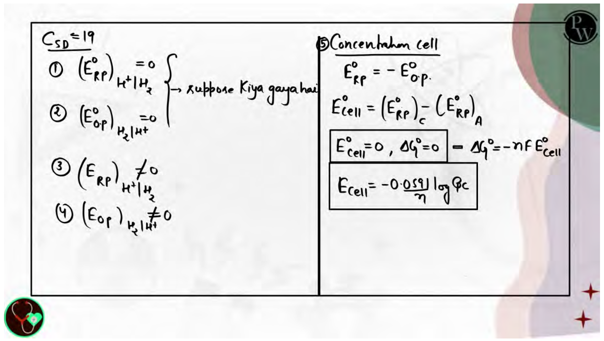
$\Delta G = -ve$

$E_{cell} = +ve$

$\log x = -ve \rightarrow x < 1$

$\log x = +ve \rightarrow x > 1$

$\log x = 0 \rightarrow x = 1$



$\underline{C_{SD} = 19}$

① $(E^\circ_{RP})_{H^+|H_2} = 0$

② $(E^\circ_{O.P})_{H_2|H^+} = 0$

} → suppose Kiya gaya hai

③ $(E_{RP})_{H^+|H_2} \neq 0$

④ $(E_{OP})_{H_2|H^+} \neq 0$

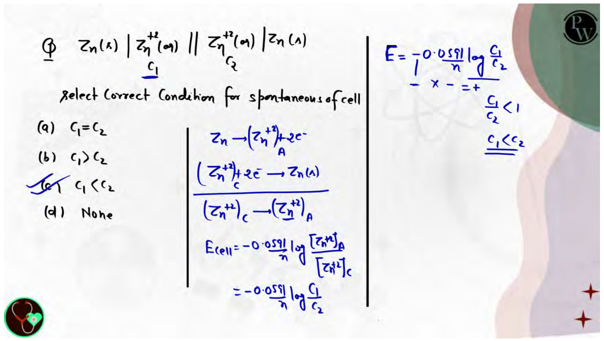
⑤ Concentration cell

$E^\circ_{RP} = -E^\circ_{O.P.}$

$E^\circ_{cell} = (E^\circ_{RP})_C - (E^\circ_{RP})_A$

$\boxed{E^\circ_{cell} = 0,\ \Delta G^\circ = 0}$ = $\Delta G^\circ = -nFE^\circ_{cell}$

$\boxed{E_{cell} = -\frac{0.0591}{n}\log Q_c}$



Q $Zn(s)\,|\,\underset{\underline{C_1}}{Zn^{+2}(aq)}\ ||\ \underset{C_2}{Zn^{+2}(aq)}\,|\,Zn(s)$

select Correct Condition for spontaneous of cell

(a) $C_1 = C_2$

(b) $C_1 > C_2$

(c) $C_1 < C_2$ ✓

(d) None

$Zn \longrightarrow (Zn^{+2})_A + 2e^-$

$(Zn^{+2})_C + 2\bar{e} \longrightarrow Zn(s)$

$(Zn^{+2})_C \longrightarrow (Zn^{+2})_A$

$E_{cell} = -\frac{0.0591}{n}\log\frac{[Zn^{+2}]_A}{[Zn^{+2}]_C}$

$= -\frac{0.0591}{n}\log\frac{C_1}{C_2}$

$E = -\frac{0.0591}{n}\log\frac{C_1}{C_2}$

$- \times - = +$

$\frac{C_1}{C_2} < 1$

$\underline{\underline{C_1 < C_2}}$

Concentration cell → $E^0_{cell}=0$, $E_{cell}\neq 0$
$\Delta G^0=0$ $\Delta G\neq 0$

Electrolyte Con. cell

Electrolyte = diff Con. in both halfcell

Electrode = same in halfcell

$E^0_{cell}=0$

$\Delta G^0=0$

Electrode Con. cell

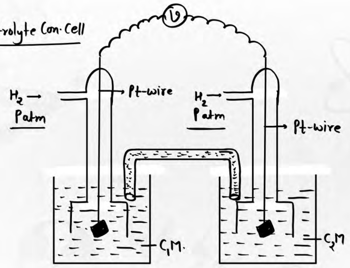
Electrode = different Con. in both halfcell

Electrolyte = same in both halfcell

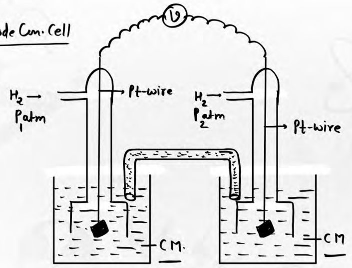
$E^0_{cell}=0$

$\Delta G^0=0$

Electrolyte Con. Cell



Electrode Con. Cell



Q Pt, $H_2(g)$ | $H^+(aq)$ || $H^+(aq)$ | $H_2(g)$, Pt is a example of
P_1 atm P_2 atm

(a) Concentration cell

(b) Electrode Con. cell

(c) Electrolyte Con. cell

(d) 'A' & 'B' Both correct

Q → select correct relation of P_1 & P_2 for spontaneous of cell

(a) $P_1=P_2$

(b) $P_1>P_2$

(c) $P_1<P_2$

(d) None

$(H_2)_A \longrightarrow 2H^+ + 2e^-$

$2H^+ + 2e^- \longrightarrow (H_2)_C$

$(H_2)_A \longrightarrow (H_2)_C$

$E=-\frac{0.0591}{n}\log\frac{P_2}{P_1}$ $\frac{P_2}{P_1}<1$

$-\ \times\ - = +v$

$P_2<P_1$

Q12. A hydrogen gas electrode is made by dipping platinum wire in a solution of HCl of pH = 10 and by passing hydrogen gas around the platinum wire at 1 atm pressure. The oxidation potential of electrode would be

$[H^+] = 10^{-10}$ $E_{O.P} = ?$ $H_2(g) \longrightarrow 2H^+(aq) + 2e^-$

[NEET 2013]

(a) 0.059 V

$\boxed{E^0_{O.P} = 0}$

[✓ (b) 0.59 V]

(c) 0.118 V

(d) 1.18 V

$$E_{O.P} = E^0_{O.P} - \frac{0.0591}{n} \log Q_c$$
$$= 0 - \frac{0.0591}{2} \log \frac{[H^+]^2}{P_{H_2}}$$
$$= -\frac{0.0591}{2} \log \frac{(10^{-10})^2}{1}$$

$$E_{O.P} = -\frac{0.0591}{\cancel{2}} \times \cancel{2} \log 10^{-10}$$
$$= E_{O.P.} = -10 \times -0.0591 \log 10$$
$$= +0.591V$$

Q19. For the reduction of silver ions with copper metal, the standard cell potential was found to be +0.46V at 25ºC. The value of standard Gibbs energy, ΔGº will be (F = 96500 C mol^{-1})

50Sec H.W.

[CBSE AIPMT 2010]

(a) –89.0 kJ

(b) –89.0 J

(c) –44.5 kJ

(d) 98.0 kJ

Q18. Standard electrode potential for Sn^{4+}/Sn^{2+} couple is +0.15V and that for the Cr^{3+}/Cr couple is –0.74. These two couples in their standard state are connected to make a cell. The cell potential will be

Cathode

50Sec

[CBSE AIPMT 2011]

[✓ (a) +0.89V]

(b) +0.18V

(c) +1.83V

(d) +1.199V

$$E^0_{Cell} = (E^0_{RP})_C - (E^0_{RP})_A$$
$$= 0.15 - (-0.74)$$
$$= 0.15 + 0.74$$
$$= +0.89$$

Q2. The pressure of H_2 required to make the potential of H_2-electrode zero in pure water at 298K is

$[H^+] = [OH^-] = 10^{-7}$ 50Sec

[CBSE AIPMT 2016, Phase I]

$E_{RP} = 0$ $2H^+ + 2e^- \longrightarrow H_2(g)$

(a) 10^{-12} atm

(b) 10^{-10} atm

(c) 10^{-4} atm

[✓ (d) 10^{-14} atm]

$$E_{RP} = E^0_{RP} - \frac{0.0591}{n} \log Q_c$$
$$0 = 0 - \frac{0.0591}{n} \log \frac{P_{H_2}}{[H^+]^2}$$
$$0 = \log \frac{P_{H_2}}{[10^{-7}]^2}$$
$$\log 1 = \log \frac{P_{H_2}}{10^{-14}}, \quad P_{H_2} = 10^{-14}$$

Q7. Zinc can be coated on iron to produced galvanised iron but the reverse is not possible. It is because

[CBSE AIPMT 2016, Phase-II]

(a) zinc is lighter than iron

(b) zinc has lower melting point than iron

(c) zinc has lower negative electrode potential than iron

(d) zinc has higher negative electrode potential than iron

Q1. In the electrochemical cell Zn || $ZnSO_4$ (0.01M) || $CuSO_4$(1.0M) Cu, the emf of this Daniel cell is E_1. When the concentration $ZnSO_4$ is changed to 1.0M and that of $CuSO_4$ changed to 0.01M, the emf changes to E_2. From the followings, which one is the relationship between E_1 and E_2? (Given, RT/F = 0.059)

[NEET 2017, 2003]

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

(a) $E_1 = E_2$

(b) $E_1 < E_2$

(c) $E_1 > E_2$ ✓

(d) $E_2 = 0 \neq E_1$

$E_1 > E_2$

$$E_{cell} = E^0_{cell} - \frac{0.0591}{n} \log Qc$$

$$E_1 = E^0 - \frac{0.0591}{2} \log \frac{Zn^{+2}}{Cu^{+2}}$$

$$E_1 = E^0 - \frac{0.0591}{2} \log \frac{0.01}{1}$$

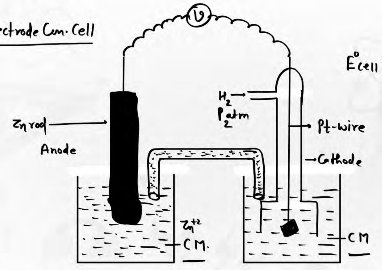
$$E_2 = E^0 - \frac{0.0591}{2} \log \frac{1}{0.01}$$

Electro-chemical Series or E.C.S. Series

Electro-chemical series. is increasing order of standard Reduction Electrode potential of halfcell with respect to hydrogen electrode

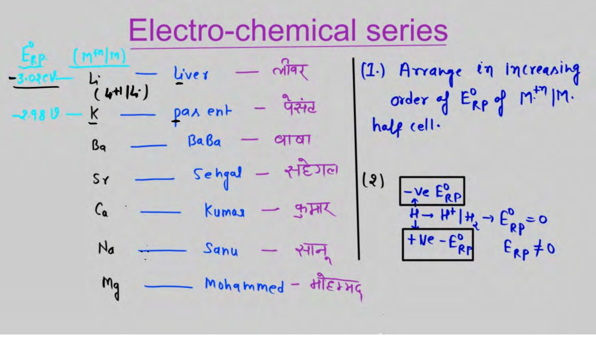
→ $E^0_{H^+|H_2} = 0$ → suppose

Electrode Con. Cell



$$E^0_{cell} = \underset{\substack{\downarrow \\ 0}}{(E^0_{RP})_C} - (E^0_{RP})_A$$

$$E^0_{cell} = (E^0_{RP})_A$$



Electro-chemical series

E^0_{RP} $(M^{+n}|M)$

$-3.02eV$ — Li — Liver — लीवर

$(Li^{+}|Li)$

$-2.98V$ — K — par ent — पैरेंट

Ba — BaBa — बाबा

Sr — Sehgal — सहेगल

Ca — Kumar — कुमार

Na — Sanu — सानू

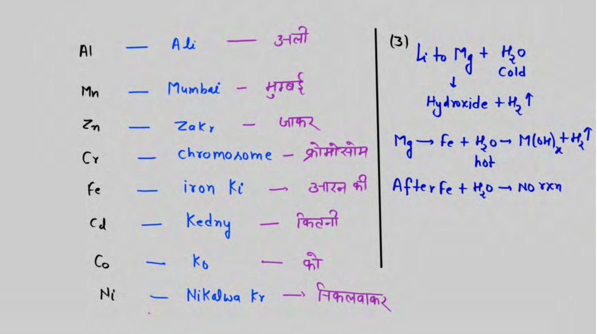
Mg — Mohammed — मोहम्मद

(1.) Arrange in increasing order of E^0_{RP} of $M^{+n}|M$ half cell.

(2) $\boxed{-ve\ E^0_{RP}}$

$H \rightarrow H^+|H_2 \rightarrow E^0_{RP} = 0$

$\boxed{+ve - E^0_{RP}}$ $E_{RP} \neq 0$



Al — Ali — अली

Mn — Mumbai — मुम्बई

Zn — Zakr — जाकर

Cr — chromosome — क्रोमोसोम

Fe — iron Ki → आयरन की

Cd — Kedny — कितनी

Co — Ko — को

Ni — Nikalwa kr → निकलवाकर

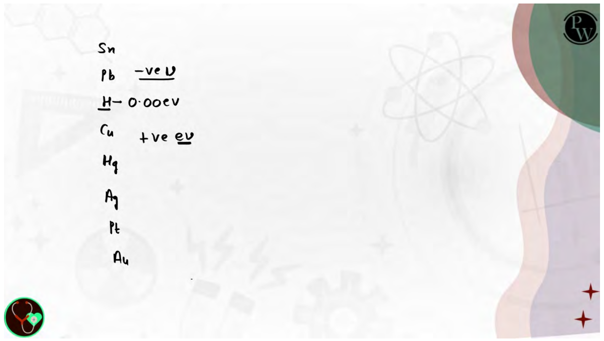
(3) Li to Mg + H_2O Cold

↓

Hydroxide + $H_2\uparrow$

$Mg \rightarrow Fe + \underset{hot}{H_2O} \rightarrow M(OH)_x + H_2\uparrow$

After Fe + $H_2O \rightarrow$ NO rxn



Sn

Pb $\underline{-ve\ V}$

$\underline{H} \rightarrow 0.00eV$

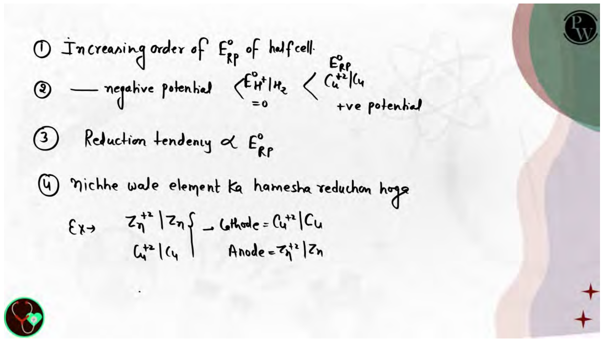
Cu +ve $\underline{eV}$

Hg

Ag

Pt

Au



① Increasing order of E^0_{RP} of half cell.

② — negative potential $< E^0_{H^+|H_2} < \overset{E^0_{RP}}{Cu^{+2}|Cu}$

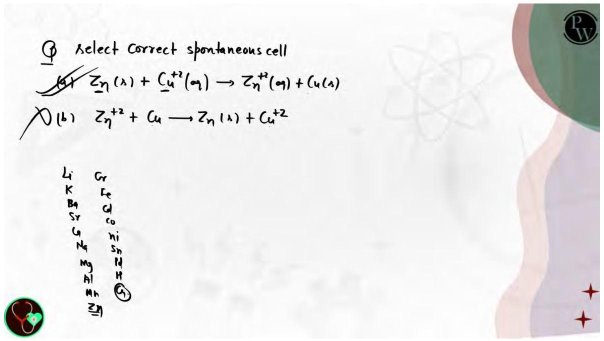
$= 0$ +ve potential

③ Reduction tendency $\propto E^0_{RP}$

④ nichhe wale element Ka hamesha reduction hoga

Ex → $\left.\begin{matrix} Zn^{+2}|Zn \\ Cu^{+2}|Cu \end{matrix}\right\}$ → Cathode = $Cu^{+2}|Cu$

Anode = $Zn^{+2}|Zn$


Q select correct spontaneous cell

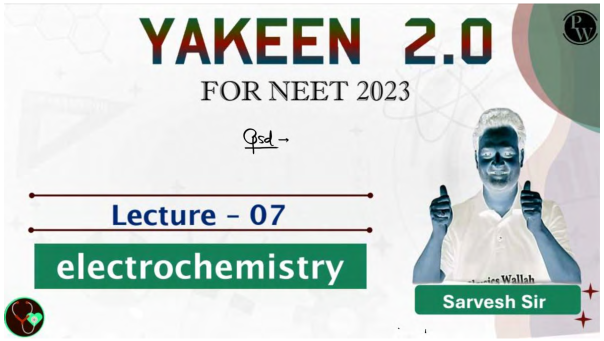
(a) $Zn(s) + Cu^{+2}(aq) \longrightarrow Zn^{+2}(aq) + Cu(s)$

(b) $Zn^{+2} + Cu \longrightarrow Zn(s) + Cu^{+2}$

Li
K
Ba
Sr
Ca
Na
Mg
Al
Mn
Zn

Cr
Fe
Cd
Co
Ni
Sn
Pb
H
Cu


THANK YOU


YAKEEN 2.0

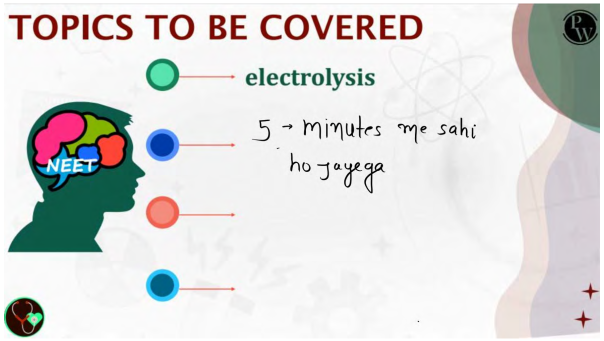
FOR NEET 2023

Qsd →

Lecture – 07

electrochemistry

Sarvesh Sir


TOPICS TO BE COVERED

- electrolysis
- 5 → minutes me sahi ho jayega

Qsd-1

the standard electrode potential of Cell is 5.91 V,

40 Sec

$\overset{n=2}{Cu + 2Ag^{+} \longrightarrow Cu^{+2} + 2Ag}$ ($n=2$)

Calculate Equilibrium Constant

(a) 10^{10}

(b) 10^{100}

(c) 10^{-100}

✓(d) None.

$$E^{\circ}_{cell} = +\frac{0.0591}{n}\log K_{eq}$$

$$\overset{100}{\cancel{5.91}} = +\frac{\cancel{0.0591}}{2}\log K_{eq}$$

$$+200 = \log K_{eq}$$

$$K_{eq} = 10^{+200}$$

Qsd-2

$A(s) \longrightarrow A(aq)^{+2}$, $E = +x\,V$, $-x\,V$

$A^{+2}(aq) \longrightarrow A^{+4}(aq)$, $E = y\,V$, $-y$

Calculate value of $A(s) \longrightarrow A^{+4}(aq)$ ④ Electrode potential

E_{op}

(a) $2x+2y$

✓(b) $\frac{x+y}{2}$

(c) $-\left(\frac{x+y}{2}\right)$

(d) None

$$E_{o.p} = -E_{RP}$$

$$= -\left(-\left(\frac{x+y}{2}\right)\right)$$

$$= +\left(\frac{x+y}{2}\right)$$

$$\boxed{E_3 = \frac{n_1E_1 + n_2E_2}{n_3}}$$

$E_1, E_2, E_3 \rightarrow E_{RP}$

$$E_{RP} = \frac{2\times(-x) + 2(-y)}{4}$$

$$E_{RP} = -\left(\frac{x+y}{2}\right)$$

Qsd-3

$Cu^{+2}(0.2M) + Zn(s) \longrightarrow Zn^{+2}(0.4) + Cu(s)$, $E^{\circ}_{cell} = E_1$

select Correct statements

(a) $E_{cell} > E_1$

✓(b) $E_{cell} < E_1$

(c) $E_{cell} = E_1$

(d) None

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0591}{n}\log Q_c$$

$$= E_1 - \frac{0.0591}{2}\log\frac{Zn^{+2}}{Cu^{+2}}$$

$$= E_1 - \frac{0.0591}{2}\log\frac{0.4}{0.2}$$

$$E_{cell} = \underline{E_1} - \frac{0.0591}{2}\underbrace{\log 2}_{+ve}$$

$E_{cell} < E_1$

Electro-chemical series

Li ($Li^{+}\vert Li$)	Liver	लीवर
K	pasent	पैसंट
Ba	BaBa	बाबा
Sr	Sehgal	सहेगल
Ca	Kumar	कुमार
Na	Sanu	सानू
Mg	Mohammed	मौहम्मद

* Arrange - increasing order of E°_{RP} of half cell.

* $(E^{\circ}_{RP})_{H^{+}|H_2} = 0$

* $(E^{\circ}_{O.P.})_{H_2|H^{+}} = 0$

→ suppose Kiya hai

$E_{RP} \neq 0$, $E_{O.P} \neq 0$ for H-halfcell

Al — Ali — अली

Mn — Mumbai - मुम्बई

Zn — Zakr - जाकर

Cr — chromosome - क्रोमोसोम

Fe — iron Ki → आरन की

Cd — Kedny — कितनी

Co — Ko — को

Ni — Nikalwa kr → निकलवाकर

$\underline{C_{SD}}$

→ Strong oxidizing agent $\propto E^{\circ}_{RP}$

→ Strong Reducing agent $\propto E^{\circ}_{O.P.}$

→ always → nichhe wala
half Reduction Karega
Sabse pahale

Sn — Tino — तीनो

Pb — log — लोग

H — High — हाई

Cu — capsule → क्रैप्सूल

Hg — Hum se lene → हम से लेने

Ag — silver ka → सिल्वर का

Pt — Paliyala → पालियाला

Au — Aaye → आए

---

1. **Given,**

$Co^{3+} + e^- \rightarrow Co^{2+}$; $E^{\circ} = + \underline{1.81}$ V

$Pb^{4+} + 2e^- \rightarrow Pb^{2+}$; $E^{\circ} = + 1.67$ V

$Ce^{4+} + e^- \rightarrow Ce^{3+}$; $E^{\circ} = + 1.61$ V

$Bi^{3+} + 3e^- \rightarrow Bi$; $E^{\circ} = + \underline{0.2}0$ V

Oxidising power of the species will increase in the order

$\propto E^{\circ}_{RP}$

(A) $Ce^{4+} < Pb^{4+} < Bi^{+3} < Co^{3+}$

✓ (B) $Bi^{3+} < Ce^{4+} < Pb^{4+} < Co^{3+}$

(C) ✗ $Co^{3+} < Ce^{4+} < Bi^{+3} < \underline{Pb}^{4+}$

(D) ✗ $Co^{3+} < Pb^{4+} < Ce^{4+} < \underline{Bi}^{3+}$

---

4. **Given, that** $E^{\ominus}_{\overset{0}{O_2}/H_2\underset{-2}{O}} = +1.23\ V$;

$E^{\ominus}_{S_2O_8^{2-}/SO_4^{2-}} = \underline{2.05 V}$;

$E^{\ominus}_{Br_2/Br^{\ominus}} = +1.09\ V$,

$E^{\ominus}_{Au^{3+}/Au} = +1.4\ V$

The strongest oxidizing agentis $\propto E^{\circ}_{RP}$

S.

(A) Au^{3+}

(B) O_2

✓ (C) $S_2O_8^{2-}$

(D) Br_2

Q16. Standard Reduction electrode potential of three metals X, Y and Z are –1.2V, +0.5V and – –3.0V respectively. The reducing power of these metals will be +1.2V –0.5V, +3.0V

[CBSE AIPMT 2011]

Reducing power $\propto E^\circ_{o.p.}$

(a) X > Y > Z

(b) X < Y < Z   Z > X > Y

(C) X < Z < Y

(d) Y < X < Z

Q26. On the basis of the following E° values, the strongest oxidising agent is $\propto E^\circ_{RP}$

[CBSE AIPMT 2008]

$\overset{+2}{[Fe(CN)_6]^{4-}} \rightarrow \overset{+3}{[Fe(CN)_6]^{3-}} + e^-$; $E^\circ = -0.35V$,

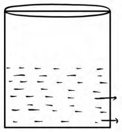
$Fe^{2+} \rightarrow Fe^{3+} + e^-$; $E^\circ = -0.77V$, –0.77V

+0.77V

(a) $[Fe(CN)_6]^{4-}$ (b) Fe^{2+}

(c) Fe^{3+} (d) $[Fe(CN)_6]^{3-}$

Q.



Mn
Zn
Fe
Cu

→ mixture

→ $Zn^{+2}(aq)$, $Fe^{+2}(aq)$, $Mn^{+2}(aq)$, $Cu^{+2}(aq)$

which ion deposite (Reduce → $M^{+n}(aq) \rightarrow M(s)$) first

(a) Zn^{+2} $Cu^{+2} > Fe^{+2} > Zn^{+2} > Mn^{+2}$

(b) Fe^{+2}

(c) Mn^{+2}

(d) Cu^{+2}

Q select the composition, In which H_2 gas release

(a) Fe + dil HCl

(b) Cu + HCl →

(c) both

(d) None

$\overset{0}{Fe} + \overset{+1}{H}Cl \longrightarrow \overset{0}{H_2} + \overset{+2}{FeCl_2} \mid \overset{+3}{FeCl_3}$

XX $\overset{0}{Cu} + \overset{+1}{H}Cl \longrightarrow \overset{0}{H_2} + \overset{+2}{CuCl_2}$

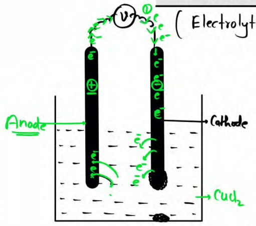
Fe

H →

Cu →

Electrolysis

(Electrolytic cell)



$Cu^{+2}(aq) + 2e^- \rightarrow Cu(s)$ – Reduct^n Cathode

$2Cl^- \rightarrow Cl_2 + 2e^-$ → Oxid^n → Anode

- ✱ Electrical E ⟶ Chemical E.
- ✱ nonspontaneous → $\Delta G > 0$
- ✱ both Electrode present in same Compartment
- ✱ Electrode → inert electrode (Pt-wire graphite rod)
- ✱ Common electrolytic solution

	Galvanic cell	Electrolyte cell
Anode ⟶	⊖	⊕
Cathode ⟶	⊕	⊖
ΔG ⟶	$=0$, $\Delta G = -ve$	→ $\Delta G > 0$
✓	$\Delta G = -nFE_{cell}$	$\Delta G = -nFE_{cell}$ → not applicable

Application

CSD

Electrolysis

Cation → Reduction

If more than one cation present in solution → Reduction tendency $\propto E^0_{RP}$

→ Jo half cell E.C.S. sabse nichhe hoga

Oxidation ⇒ Anion

$X^- > OH^- > NO_3^- > SO_4^{-2}$

$2X^- \longrightarrow X_2\uparrow + 2e^-$

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

Ex → aq NaCl solution → Electrolysis

Cation → Na^+, H^+ — $2H^+ + 2e^- \rightarrow H_2(g)$ → Cathode

Anion → Cl^-, OH^- ⟶ $2Cl^- \longrightarrow Cl_2 + 2e^-$ ⟶ Anode

Cell Rxn = $\overset{+1}{2H^+} + \overset{-1}{2Cl^-} \longrightarrow \overset{0}{H_2}\uparrow + \overset{0}{Cl_2}\uparrow$

Red = Cathode

Oxd = Anode

Ex → aq $CuSO_4$ solution

Cation → Cu^{+2}, H^+ → $Cu^{+2} + 2e^- \longrightarrow Cu(s)$ → Cathode

Anion → SO_4^{-2}, OH^- → $4OH^- \longrightarrow O_2 + 2H_2O + 4e^-$ → Anode

$$\underset{+2}{2Cu^{+2}} + 4\overset{-2}{O}H^- \longrightarrow \uparrow \overset{0}{O_2} + 2H_2O + \underset{0}{2Cu(s)}$$

oxdⁿ = Anode

Red = Cathode

Q In electrolysis of aqueous solution of $CuSO_4$.

Q-1 Write Anode Reaction → $4OH^- \longrightarrow \uparrow O_2 + 2H_2O + 4e^-$

Q-2 Write Cathode Reaction → $Cu^{+2} + 2e^- \longrightarrow Cu(s)$

Q-3 Write Cell Rxn —— $2Cu^{+2} + 4OH^- \longrightarrow 2Cu_{(s)} + O_2\uparrow + 2H_2O(l)$

Q-4 which gas release in electrolysis → $O_2\uparrow$

Q-5 which gas release at Anode in electrolysis → $O_2\uparrow$

Q-5 → pH of solution increases|decreases After Electrolysis → pH decrease

$p^{OH} = -\log[OH^-]$

$\underset{\downarrow \text{ increase}}{p^{OH}} = \log \dfrac{1}{\underset{\downarrow \text{ decreases}}{[OH]}}$

$\underset{\downarrow \text{ decreasing}}{pH} + \underset{\downarrow \text{ increases}}{p^{OH}} = 14$

Q-6 → what amount of Cu-deposite

Q-7 → what vol. of O_2 gas release at S.T.P. ↑ = ?

} → Faraday's Law

Q-8 → what Concentration of Na^+ change After electrolysis of aq. solⁿ of Na_2SO_4

(a) increases

(b) decreases

[✓ (c)] Remains same

(d) 1st decreases than increases

Cation → H^+, Na^+ → $2H^+ + 2e^- \rightarrow H_2\uparrow$ → Cathode

Anion → SO_4^{-2}, OH^- → $4OH^- \rightarrow \uparrow O_2 + 2H_2O + 4e^-$ → Anode

$$4H^+ + 4OH^- \longrightarrow H_2 + O_2 + 2H_2O$$

Q select correct Cathode rxn of molten or fused NaCl

(a) $2H^+ \longrightarrow H_2$

[✓ (b)] $Na^+ \rightarrow Na$

(c) $Cl_2 \longrightarrow 2Cl^-$

(d) None

molten or fused → absence of H_2O

Cation → Na^+ → $Na^+ + e^- \rightarrow Na(s)$

Anion → Cl^- → $2Cl^- \longrightarrow Cl_2 + 2e^-$

$$2Na^+ + 2Cl^- \longrightarrow 2Na(s) + Cl_2(g)\uparrow$$

Faraday's Law

1st law ⇒ total amount deposite or total gas release is directly proportional to Actual Charge utilized in electrolysis

$w \propto q$

$w = Zq$

↳ electro-chemical constant

$\boxed{w = Z \times i \times t}$ $\left(q = \underset{\text{Currect}}{i} \times \underset{\text{time}}{t}\right)$

CSD

$q = 1F = 96500\,C$ Charge = Charge of 1 mole e^-

$w =$ 1 gram Equivalent metal deposite hoga or gas release hogi
or
Equivalent wt. in gram metal deposite

1 gram Equivalent Contains mass = E.wt in gram
1 mole Contains wt = M.wt in gram

No. of gram Eq $= \frac{wt}{E.wt}$

$1 = \frac{wt}{E.wt}$

$wt = (E.wt)$ gram

Q $CuSO_4(aq)$
1F Charge
1 gram Equivalent of Cu-deposite
= 31.75 gram Cu-deposite

$E.wt \text{ of } Cu = \frac{At.wt}{\eta\text{-factor}} = \frac{63.5}{2} = 31.75$ gram

η-factor of Metal ion

= charge of metal

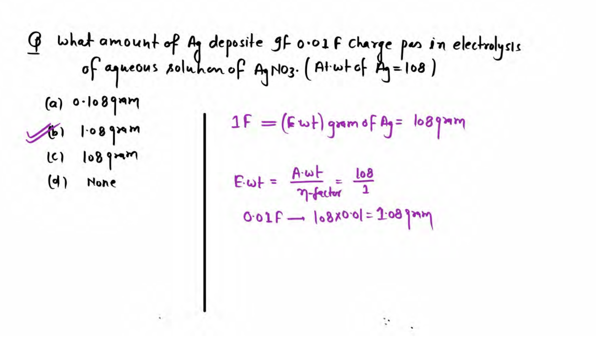
$Na^{+} \rightarrow \eta = 1$, $Cr^{+3} \rightarrow \eta = 3$
$Cu^{+2} \rightarrow \eta = 2$ $Zr^{+4} \rightarrow \eta = 4$
$Zn^{+2} \rightarrow \eta = 2$
$Al^{+3} \rightarrow \eta = 3$

Q What amount of Zn deposite at Cathode if 1F Charge pas in electrolysis of molten $ZnSO_4$
(At.wt of Zn = 65)

(a) 65 gram
(b) $\frac{65}{2}$ gram
(c) 1 gram Equivalent
[✓ (d)] 'B' & 'C' both Correct

1F = 1 gram Equivalent
= (E.wt) gram Zn deposite

$E.wt \text{ of } Zn = \frac{A.wt}{\eta\text{-factor}} = \frac{65}{2}$



Q what amount of Ag deposite gf 0·01 F charge pas in electrolysis of aqueous soluhon of $AgNO_3$. (At·wt of Ag = 108)

(a) 0·108 gram

[✓ (b)] 1·08 gram

(c) 108 gram

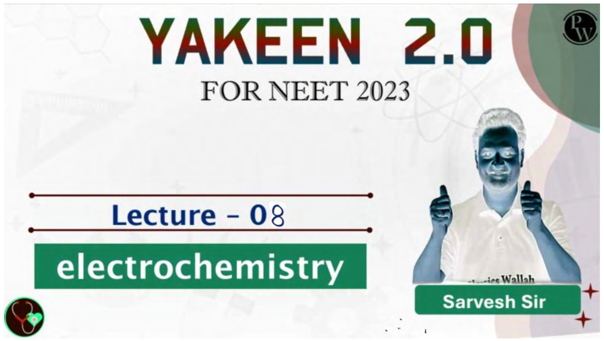
(d) None

1F = (E wt) gram of Ag = 108 gram

$$E \cdot wt = \frac{A \cdot wt}{\eta\text{-factor}} = \frac{108}{1}$$

$0 \cdot 01F \longrightarrow 108 \times 0 \cdot 01 = 1 \cdot 08$ gram

THANK YOU



YAKEEN 2.0

FOR NEET 2023

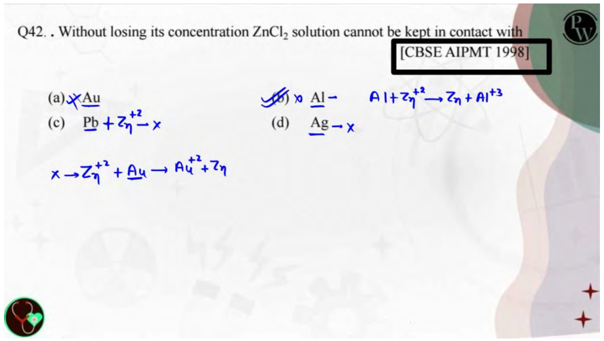
Lecture – 08

electrochemistry

Sarvesh Sir

Today's Goal

- ★ Faraday's law
- ★ Battery



Q42. . Without losing its concentration $ZnCl_2$ solution cannot be kept in contact with

[CBSE AIPMT 1998]

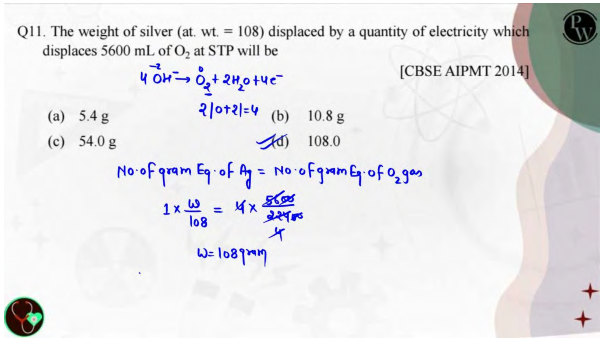
(a) ✗ $\underline{Au}$

✓(b) ✗ $\underline{Al} \rightarrow$ $Al + Zn^{+2} \longrightarrow Zn + Al^{+3}$

(c) $\underline{Pb} + Zn^{+2} \rightarrow$ ✗

(d) $\underline{Ag} \rightarrow$ ✗

$\times \rightarrow Zn^{+2} + \underline{Au} \longrightarrow Au^{+2} + Zn$



Q11. The weight of silver (at. wt. = 108) displaced by a quantity of electricity which displaces 5600 mL of O_2 at STP will be

[CBSE AIPMT 2014]

$4\overset{-2}{O}H^- \rightarrow \overset{0}{O_2} + 2H_2O + 4e^-$

$2|0+2| = 4$

(a) 5.4 g

(b) 10.8 g

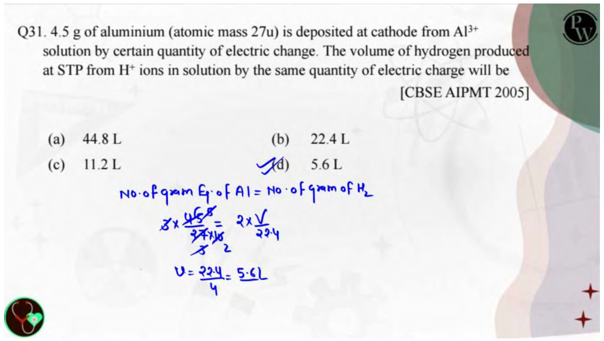
(c) 54.0 g

✓(d) 108.0

No. of gram Eq. of Ag = No. of gram Eq. of O_2 gas

$1 \times \frac{w}{108} = \cancel{4} \times \frac{\cancel{5600}}{\underset{\cancel{4}}{\cancel{22400}}}$

$W = 108$ gram



Q31. 4.5 g of aluminium (atomic mass 27u) is deposited at cathode from Al^{3+} solution by certain quantity of electric change. The volume of hydrogen produced at STP from H^+ ions in solution by the same quantity of electric charge will be

[CBSE AIPMT 2005]

(a) 44.8 L

(b) 22.4 L

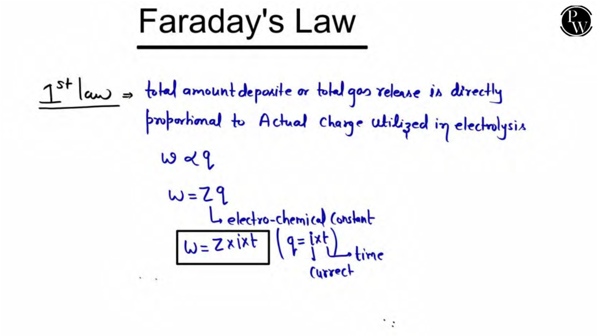
(c) 11.2 L

✓(d) 5.6 L

No. of gram Eq. of Al = No. of gram of H_2

$\cancel{3} \times \frac{\cancel{45}}{\underset{\cancel{3} \quad 2}{\cancel{27} \times \cancel{10}}} = 2 \times \frac{V}{22.4}$

$V = \frac{22.4}{4} = \underline{5.6L}$



Faraday's Law

$\underline{1^{st}\text{ law}}$ ⇒ total amount deposite or total gas release is directly proportional to Actual Charge utilized in electrolysis

$w \propto q$

$w = Zq$

↳ electro-chemical constant

$\boxed{w = Z \times i \times t}$ $(q = i \times t)$ ↳ time, current

CSD

$q = 1F = 96500C$ charge = Charge of 1mole e^-

$W =$ 1gram Equivalent metal deposite hoga or gas release hogi

or

Equivalent wt. in gram metal deposite

1gram Equivalent Contains mass = E·wt ingram | No·of gram Eq = $\frac{wt}{E \cdot wt}$

1mole Contains wt = m·wt ingram

$1 = \frac{wt}{E \cdot wt}$

$wt = (E \cdot wt)$ gram

Q $CuSO_4$ (aq)

1F Charge

1 gram Equivalent of Cu-deposite

= 31·75 gram Cu-deposite

E·wt of Cu = $\frac{At \cdot wt}{\eta\text{-factor}} = \frac{63 \cdot 5}{2} = 31 \cdot 75$ gram

n-factor of Metal ion

= charge of metal

$Na^+ \rightarrow \eta = 1$, $Cr^{+3} \rightarrow \eta = 3$

$Cu^{+2} \rightarrow \eta = 2$ $Zr^{+4} \rightarrow \eta = 4$

$Zn^{+2} \rightarrow \eta = 2$

$Al^{+3} \rightarrow \eta = 3$

Q What amount of Zn deposite at Cathode if 1F charge pas in electrolysis of molten $ZnSO_4$

(At·wt of Zn = 65)

(a) 65 gram

(b) $\frac{65}{2}$ gram

(c) 1 gram Equivalent

[✓ (d) 'B' & 'C' both Correct]

1F = 1gram Equivalent

= (E·wt) gram Zn deposite

E·wt of Zn = $\frac{A \cdot wt}{\eta\text{-factor}} = \frac{65}{2}$

Q What amount of Ag deposite if 0·01F charge pas in electrolysis of aqueous solution of $AgNO_3$. (At·wt of Ag = 108)

(a) 0·108 gram

[✓ (b) 1·08 gram]

(c) 108 gram

(d) None

1F = (E·wt) gram of Ag = 108 gram

E·wt = $\frac{A \cdot wt}{\eta\text{-factor}} = \frac{108}{1}$

0·01F ⟶ 108×0·01 = 1·08 gram

$q = 96500\,C = 1F$

$W =$ E·wt gram mass deposite

$W = Z \times q$

$E{\cdot}wt = Z \times 96500$

$$\boxed{Z = \frac{E{\cdot}wt}{96500}}$$

Value of z put in $W = Z \times q$ equation

$$W = \frac{E{\cdot}wt}{96500} \times q$$

$$\frac{W}{E{\cdot}wt} = \frac{q}{96500}$$

$$\boxed{\text{No·of gram Equivalent mass deposite} = \frac{q}{96500}}$$

$\underline{C_{SD}}$

$$\text{No·of gram Eq· mass deposite | gas release} = \frac{q}{96500}$$

$$\rightarrow \boxed{\eta \times \text{No·of mole deposite | gas release} = \frac{q}{96500} = \frac{i_{actual} \times t}{96500}}$$

$$\boxed{\eta \times \frac{w}{A{\cdot}wt} = \frac{q}{96500}}$$

$$\boxed{\eta \times \frac{\text{Vol·of gas at S.T.P (L)}}{22.4} = \frac{q}{96500}}$$

→ η-factor of metal ion = charge of metal

$Zn^{+2} \rightarrow \eta = 2$, $Zr^{+4} \rightarrow \eta = 4$, $Al^{+3} \rightarrow \eta = 3$

Ex → $2\overset{-1}{X^-} \longrightarrow \uparrow \overset{0}{X}_{\underline{2}} + 2e^- \rightarrow 2|0+1| = \underline{2}$, $X_2 \rightarrow \eta = 2$

Ex → $4\overset{-2}{OH^-} \longrightarrow \overset{0}{O}_2 + 2H_2O + 4e^- \rightarrow 2|0+2| = 4 \rightarrow O_2 \rightarrow \eta = 4$

$$\rightarrow \text{Current efficiency } (\eta) = \frac{i_{actual}}{i_{theoretical}} \times 100$$

$$\boxed{\rightarrow i_{actual} = \frac{\text{current efficiency} \times i_{theoretical}}{100}}$$

---

2. **A solution of $\underline{\overset{+2}{Ni}}(NO_3)_2$ is electrolysed between platinum electrodes using 0.1 Faraday electricity. How many mole of Ni will be deposited at the cathode?**

$Ni = 58.5$

(A) **0.20**

(B) **0.10**

(C) **0.15**

[✓ **(D)** **0.05**]

q

α,

$$W = \frac{58.5 \times 0.1}{2} =$$

$$\eta \times \text{mole} = \frac{q}{96500}$$

$$2 \times \text{mole} = \frac{0.1F}{1F}$$

$$\text{Mole} = \frac{0.1}{2} = \underline{\underline{0.05}}$$

10. Two Faraday of electricity is passed through a solution of $\underset{+2}{Cu}SO_4$. The mass of copper deposited at the cathode is
(at. mass of Cu = $\underline{63.5}$ u)

(A) 0 g
[✓ (B) 63.5 g]
(C) 2 g
(D) 127 g

$\eta \times mole = \frac{q}{96500}$

$\cancel{2} \times \frac{W}{63.5} = \frac{\cancel{2}}{1F}$

W = 63.5 gram

12. Electrolysis of dilute aqueous NaCl solution was carried out by passing 10 mA current. The time required to liberate 0.01 mole of H_2 gas at the cathode is (1 F = 96500 C mol^{-1})

$2\overset{+1}{H^+} + 2e^- \rightarrow \overset{0}{H_{\underline{2}}}$

$2|0-1| = 2$

(A) 9.65×10^4 s
[✓ (B) 19.3×10^4 s]
(C) 28.95×10^4 s
(D) 38.6×10^4 s

$\eta \times mole = \frac{q}{96500}$

$2 \times 0.01 = \frac{10 \times 10^{-3} \times t}{96500}$

$\frac{2 \times 0.01 \times 96500}{10 \times 10^{-3}} = t$

$t = \frac{2 \times 965}{10^{-2}}$

$t = 1930 \times 10^2$

$= \underline{19.3 \times 10^4 \text{ sec}}$

21. The electric charge for electrode deposition of one gram equivalent of a substance is

(A) one ampere per second
[✓ (B) 96,500 coulombs per second]
(C) one ampere for one hour
(D) charge on one mole of electrons

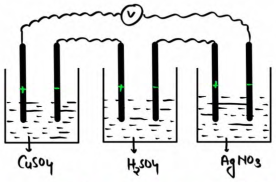
23. Faraday's laws of electrolysis are related to the

(A) atomic number of the cation
(B) atomic number of the anion
[✓ (C) equivalent weight of the electrolyte]
(D) speed of the cation

2^{nd} law of faraday's

q = charge ⇒ same utilized

No. of gram Eq deposite $= \frac{q}{96500}$



"If No. of cell arrange in series than same No. of gram Equivalent deposite in each compartment"

No. of gram Eq. of Cu-deposite = No. of gram Eq. of Ag-deposite = No. of gram Eq of H_2 gas releases

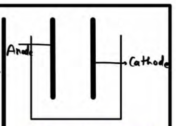
---

CSD

If No. of cell arrange in series than same gram Equivalent deposite in each cell

A cell + B-cell + C-Cell

$$\boxed{\frac{W_A}{(E.wt)_A} = \frac{W_B}{(E.wt)_B} = \frac{W_C}{(E.wt)_C}}$$



No. of gram Eq. of Anode = No. of gram Eq. of Cathode

$$\left(\frac{W}{E.wt}\right)_A = \left(\frac{W}{E.wt}\right)_C$$

Q In electrolysis of aq. sol^n. of NaCl, 2.24L H_2 gas release at Anode at S.T.P. How many litre Cl_2 gas release at Cathode at S.T.P.

$2\bar{e}+2H^+ \longrightarrow H_2$

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

[✓ (a)] 2.24L
(b) 1.12L
(c) 4.48L
(d) None

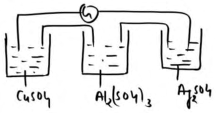
No. of gram Eq of Anode = No. of gram Eq. of Cathode

$(n \times mole)_A = (n \times mole)_C$

$\cancel{2}\times\frac{2.24}{\cancel{22.4}} = \cancel{2}\times\frac{V}{\cancel{22.4}}$

$V = 2.24L$

Q.



If equal charge utilized in electrolysis than findout deposite molar ratio of Cu, Al & Ag.

3 : 2 : 6 → Ans

gram Eq. of Cu = gram Eq. of Al = gram Eq. of Ag

$2 \times (mole)_{Cu} = 3 \times (mole)_{Al} = 1 \times (mole)_{Ag} = x$

$(mole)_{Cu} = \frac{x}{2}$
$(mole)_{Al} = \frac{x}{3}$
$(mole)_{Ag} = x$

$\frac{x}{2} : \frac{x}{3} : x$

$\frac{1}{2} : \frac{1}{3} : 1$

$\frac{3 : 2 : 6}{6}$

Battery

Combination of cell, when arrange in series, is known as battery.

Battery

primary battery

- → nonreachargable
- → Chemical E → Electrical E.

Secondary battery

- → Reachargable
- → C.E $\rightleftharpoons$ E.E

Example of primary battery

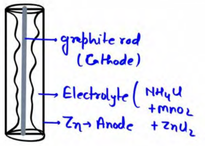
(a) Dry cell or La-lenche cell

Anode → $\overset{0}{Zn}(s) \longrightarrow \overset{+2}{Zn}^{+2} + 2e^-$

Cathode → $\overset{+4}{Mn}O_2 + NH_4^+ \longrightarrow \overset{+3}{Mn}(OH)O + NH_3\uparrow$

→ $E_{cell} = +1.51V$

→ during use, potential of cell decreases



(b) Batton cell or Zn-Hg cell

Anode → Zn-Hg →

Cathode → HgO paste

Electrolyte → KOH

Cell Reaction → $\underset{(s)}{\overset{0}{Zn}\text{-}Hg} + \underset{(l)}{\overset{+2}{Hg}O} \longrightarrow \underset{(s)}{\overset{+2}{Zn}O} + \overset{0}{Hg}(l)$

Ecell = +1.31 = Constant → during use does not change

$E = E^0 - \frac{0.0591}{n} \log \frac{Qc}{1}$

$\underline{E = E^0}$

Example of Secondary battery

Lead-battery

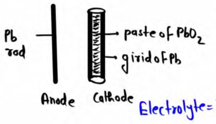
Anode → $\overset{0}{Pb} + H_2SO_4 \longrightarrow \overset{+2}{Pb}SO_4 + 2H^+ + 2e^-$

Cathode → $\overset{+4}{Pb}O_2 + H_2SO_4 \longrightarrow \overset{+2}{Pb}SO_4 + SO_4^{-2} + 2H_2O$

$\underset{(aq)}{2H_2SO_4} + \underset{(s)}{Pb} + \underset{(s)}{PbO_2} \longrightarrow \underset{(s)}{PbSO_4} + \underset{(s)}{PbSO_4} +$

$E_{cell} = +2.041V$

during process decreasing



Electrolyte = 38% (w/w) H_2SO_4 sol^n

discharge Rxn → C.E → E.E. ⇒ $Pb + PbO_2 \longrightarrow PbSO_4 + PbSO_4$

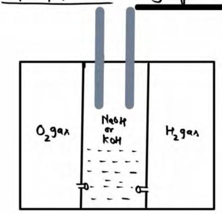
Charge Rxn → E.E. → C.E. ⇒ $PbSO_4 + PbSO_4 \rightarrow Pb + PbO_2$

Fuelcell

Combustion → $CH_4 + O_2 \longrightarrow CO_2 + H_2O$, $\Delta H = -ve$ (Enthalpy of rxn)
↓
fuel cell Convert heat of (thermal E.) rxn direct into electrical energy

Heat of rxn → direct Convert Krta hai electrical E.

Example of fuelcell → Hydrogen-oxygen fuel cell



Cell → $2\overset{0}{H_2} + \overset{0}{O_2} \rightleftharpoons 2\overset{+1}{H_2}\overset{-2}{O}$, $\Delta H = -ve$ highly exothermic

oxd, Red

efficiency of fuel cell % $= \frac{\Delta H}{\Delta G} \times 100$

$= \frac{\Delta H}{-nFE_{cell}} \times 100$

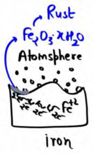
$\eta = \frac{\Delta H}{\Delta G}$

Corrosion

Natural tendency of metal to Convert its mineral Comp. (oxide, CO_3^{-2} SO_4^{-2} etc) in presence of atomsphere.

→ spontaneous process

Example → Rusting of iron



$Fe(s) \longrightarrow Fe^{+2} + 2e^-$ → Anode

$\overset{0}{O_2} + 2e^- + 2H_2O \longrightarrow 4\overset{-2}{O}H^-$ → Cathode

and form FeO, FeO further oxidise in presence of atomsphere and form $\underset{+3}{Fe_2}O_3 \cdot xH_2O$ (Rust)

Q8. A device that converts energy of combustion of fuels like hydrogen and methane, directly into electrical energy is known as

[CBSE AIPMT 2015]

(a) fuel cell
(b) electrolytic cell
(c) dynamo
(d) Ni-Cd cell

H.W.

Q27. The efficiency of a fuel cell is given by

[CBSE AIPMT 2007]

(a) $\Delta G/\Delta S$
(b) $\Delta G/\Delta H$
(c) $\Delta S/\square G$
(d) $\Delta H/\Delta G$

H.W.

HOME WORK

→ Complete all DPP
– pw modules
→ PYQ

THANK YOU