YAKEEN NEET 2.0

2026

Electrochemistry

Physical Chemistry | Lecture -6

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Topics to be covered

1. Medics Test, Revision of Last Class
2. NERNST EQUATION
3. Concentration cells
4. Home work from modules

Rule to Attend Class

1. Always sit in a peaceful environment with headphone and be ready with your copy and pen.
2. Never ever attend a class from in between or don't join a live class in the middle of the chapter.
3. Make sure to revise the last class before attending the next class & always complete your home work along with DPP.
4. Never ever engage in chat whether live or recorded on the topic which is not being discussed in current class as by doing so u can be blocked by the admin team or your subscription can be cancelled.

Rule to Attend Class

5. Try to make maximum notes during the class if something is left then u can use the notes pdf after the class to complete the remaining class.
6. Always ask your doubts in doubt section to get answer from faculty. Before asking any doubt please check whether same doubt has been asked by someone or not.
7. Don’t watch the videos in high speed if you want to understand better.


There is one big flaw in your Preparation that's name is Backlog ? What do we say to Backlog ?
NOT TODAY !!!

MEDICS

Mastery
Checks your grasp over NEET-level concepts

Evaluation
Judging both knowledge and test-smartness

Decision Making
Testing your speed + accuracy under pressure

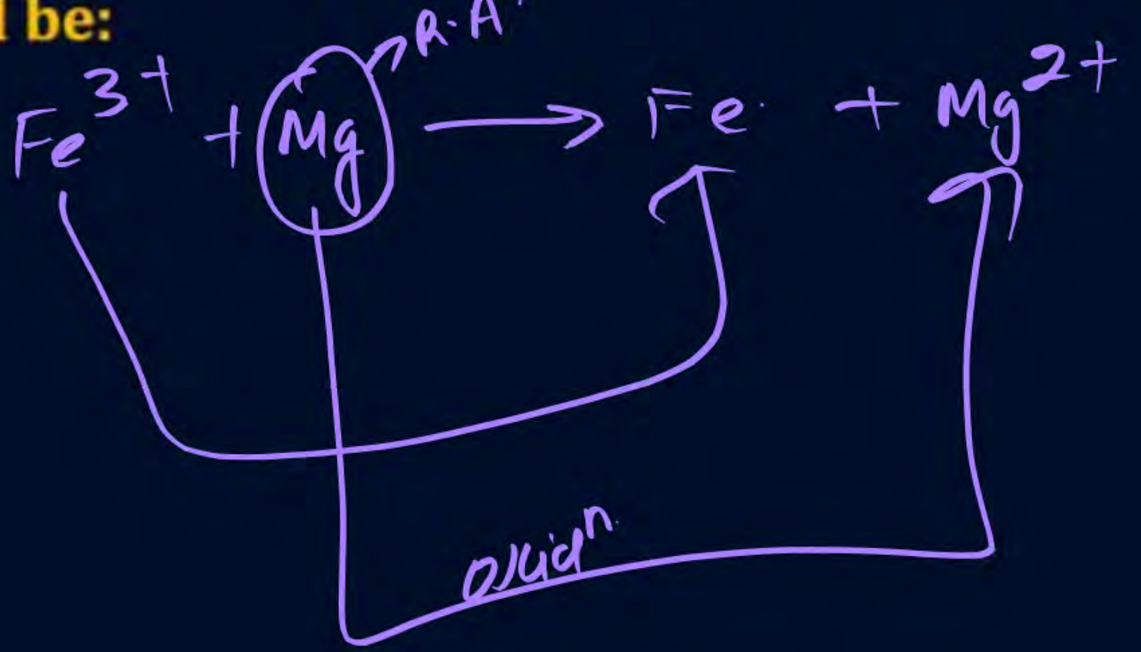
Intuition
Some answers need gut + logic – can you spot the trick?

Concepts
It's all about strong basics – no shortcuts here

Strategy
The MEDICS test – built for those who heal, hustle, and hope.

Question

Consider the cell potentials $E^{o}_{Mg^{2+}/Mg}$ = −2.37V and $E^{o}_{Fe^{3+}/Fe}$ = −0.04. The best reducing agent would be:



A Mg^{2+}

B Fe^{3+}

C Mg

D Fe

Question

The $E^{o}_{M^{3+}/M^{2+}}$ values for Cr, Mn, Fe and Co are −0.41, +1.57, 0.77 and +1.97 V respectively. For which one of these metals the change in oxidation state from +2 to +3 is easiest?

$M^{2+} \longrightarrow M^{3+} + 1e^{-}$

SRP ↓ or S.O.P ↑ ⇒ oxidn

A Cr

B Mn

C Fe

D Co

Question

Which of the following facts is not true?

M → S.R.P = (−)ve
Mg
H → S.R.P = 0
Cu
M → S.R.P = −(+)ve

$M^0 + H^{+1}Cl \rightarrow M^{+1}Cl + H_2^0$

(A) If $E^\circ(M^{n+}/M)$ is negative, H^+ will be reduced to H_2 by the metal M.

$H_2 + M^{+1}Cl \rightarrow HCl + M^0$

(B) If $E^\circ(M^{n+}/M)$ is positive, M^{n+} will be reduced to M by H_2.

(C) In a cell M^{n+}/M electrode is attached to hydrogen-half cell. To produce spontaneous cell reaction, metal M generally acts as negative electrode.

$E^\circ_{cell} = (+)ve$

(D) Compounds of active metals (Zn, Na, Mg) are reducible by H_2 whereas those of noble metals (Cu, Ag, Au) are not reducible.

$\overset{+2}{Zn}O + Hg \rightarrow \overset{0}{Hg} + \overset{0}{Zn}$

Zn

Hg → Reduce

Question

Adding powdered Pb and Fe to a solution containing 1.0 M each of Pb^{2+} and Fe^{2+} ions would result into the formation of :

$Fe^{2+} \rightarrow$ oxi

$Pb^{2+} \rightarrow$ S.R.P.↑ ∴ reduce

$Pb^{2+} + Fe \rightarrow Fe^{2+} + Pb$

(A) more of Pb and Fe^{2+} ions

(B) more of Fe and Pb^{2+} ions

(C) more of Fe and Pb

(D) more of Fe^{2+} and Pb^{2+} ions

Question

T = 298 K
C = 1 M

H_2 → anode
Cu → Reduce → Cathode

Which cell will measure standard electrode potential of copper electrode?

A. $Pt(s) \mid H_2(g, 0.1\ bar) \mid H^+(aq., 1\ M) \parallel Cu^{2+}(aq., 1\ M) \mid Cu$

B. $Pt(s) \mid H_2(g, 1\ bar) \mid H^+(aq., 1\ M) \parallel Cu^{2+}(aq., 2\ M) \mid Cu$

C. $Pt(s) \mid H_2(g, 1\ bar) \mid H^+(aq., 1\ M) \parallel Cu^{2+}(aq., 1\ M) \mid Cu$

D. $Pt(s) \mid H_2(g, 1\ bar) \mid H^+(aq., 0.1\ M) \parallel Cu^{2+}(aq., 1\ M) \mid Cu$

Question

$Sn^{4+}(aq) + 2e^- \longrightarrow Sn^{2+}(aq)$ $\quad E^\circ = 0.15V$

$Cr^{3+}(aq) + e^- \longrightarrow Cr^{2+}(aq)$ $\quad E^\circ = -0.41V$

$E^\circ_{cell} = 0.15 - (-0.41)$

$= +0.56\ V$

According to the standard reduction potentials above, what is the value of E° for the reaction below?

$$2Cr^{3+}(aq) + Sn^{2+}(aq) \longrightarrow 2Cr^{2+}(aq) + Sn^{4+}(aq)$$

Cathode Anode

$E^\circ = -0.41 - 0.15 = -0.56\ V$

A. −0.97 V

B. −0.56 V ✓

C. +0.56 V

D. +0.97 V

Electrochemistry test

↓

Lec - 1 to 6 Complete

Revision of Last Class

Nernst eqⁿ :-

↓

Cf / M ion at T ≠ 298 K

4 = 5 - 1

3 = 5 - 2

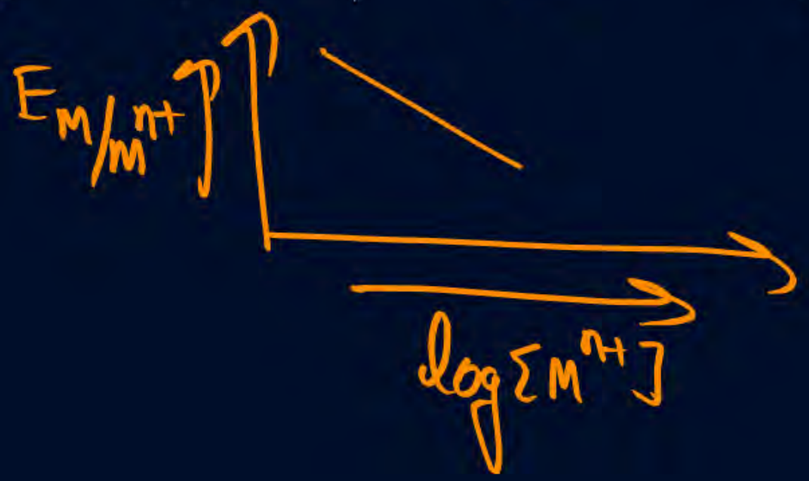
$M(s) \rightarrow M^{n+}(c_1) + ne^-$

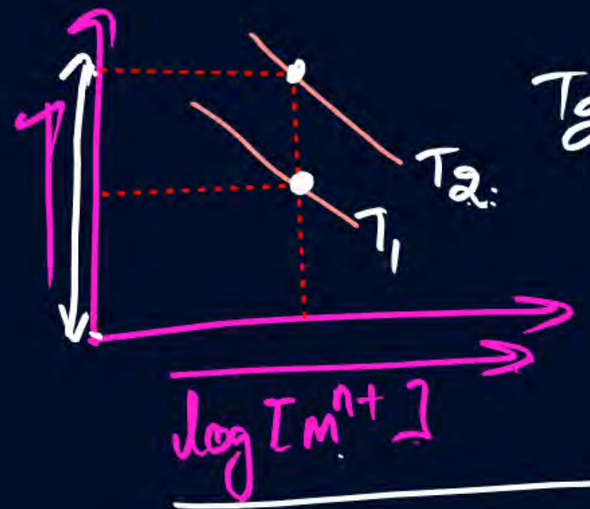
$E_{M/M^{n+}} = E^{\circ}_{M/M^{n+}} - \frac{2.303RT}{nF}\log[M^{n+}]$

$E_{M/M^{n+}} = C - m\log[M^{n+}]$

$y = C + mx$

$T \uparrow \quad E_{M/M^{n+}} \downarrow$





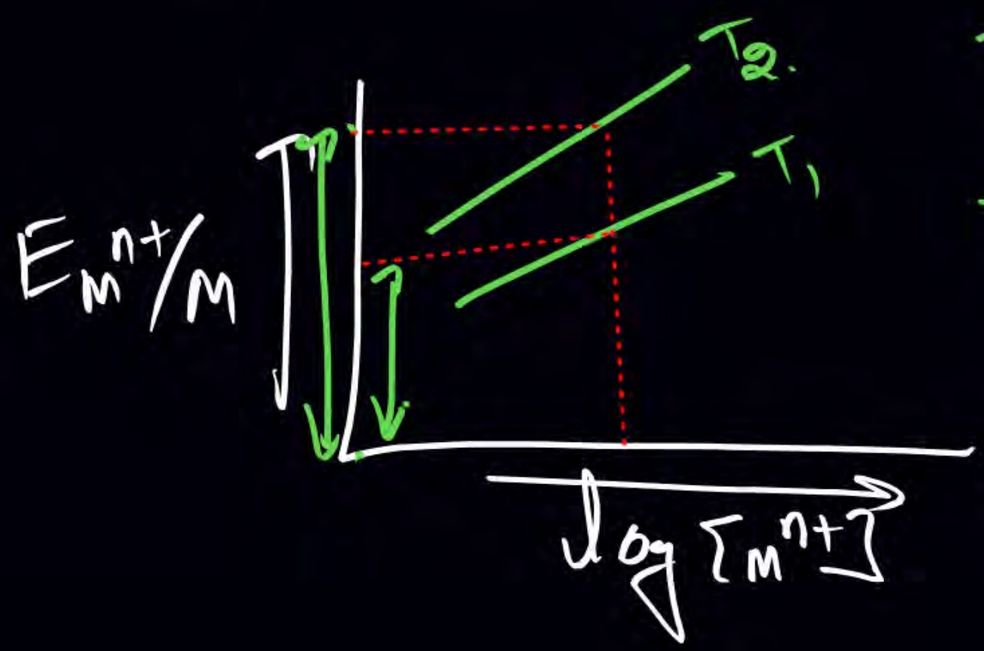
$T_2 < T_1$

$$E_{M^{n+}/M} = E^{\circ}_{M^{n+}/M} + \frac{2.303RT}{nF} \log [M^{n+}]$$

$$y = c + m x$$

6 = 5 + 1

7 = 5 + 2



$T \uparrow E_{M^{n+}/M} \uparrow$

$T_2 > T_1$

$$E_{cell} = E^{\circ}_{cell} - \frac{2.303RT}{nF} \log \frac{[P]^x}{[R]^y}$$

$$\frac{2.303RT}{F} = 0.059 \approx 0.06$$

$T = 298\ K.$

ElectroChemistry → Revise
+
Ionic eq → Revise
↓
questions

Question

$Cu^{2+} + 2e^- \rightarrow Cu$

$n = 2$

pOH = 0

$[OH^-] = 1\ M$

$E^o_{Cu^{2+}/Cu}$ = 0.34V. Calculate reduction potential at pH = 14 for the Cu^{2+}/Cu couple. K_{sp} of $Cu(OH)_2 = 1 \times 10^{-19}$.

pH = 14

$K_{sp} = [Cu^{2+}][OH^-]^2$

$10^{-19} = [Cu^{2+}](1)^2$

$[Cu^{2+}] = 10^{-19}$

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

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

Ans $E_{Cu^{2+}/Cu} = E^o_{Cu^{2+}/Cu} + \frac{0.06}{n} \log [Cu^{2+}]$

$= 0.34 - 19 \times \frac{0.06}{2} \times \log 10$

$= 0.34 - 0.57$

$= -0.23\ V$

Question

$n = 1$

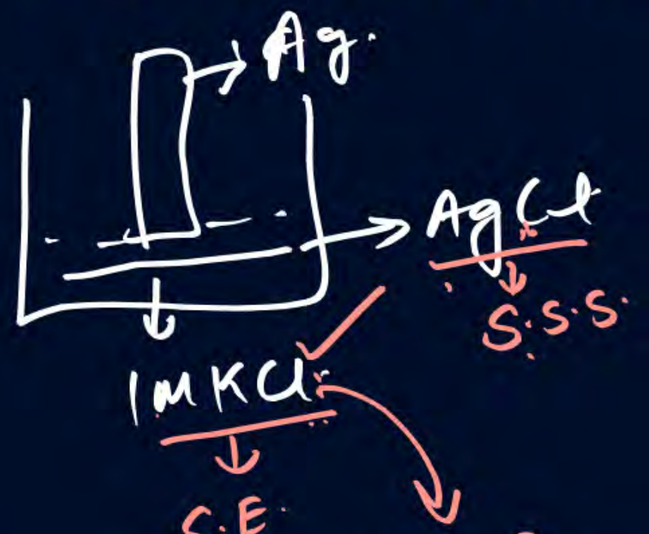
Knowing that, K_{sp} for AgCl is 1.0×10^{-10}, calculate E for Ag^+/Ag electrode immersed in 1.00 M KCl at 25°C. ($E^o_{Ag^+/Ag} = 0.799$ V)

$$E_{Ag^+/Ag} = E^o_{Ag^+/Ag} + \frac{0.06}{1} \log [Ag^+] \quad (-10)$$

$$= 0.799 - 10 \times 0.06 \log 10$$

$$= 0.799 - 0.6$$

$$= 0.199 \text{ V}$$



$$K_{sp} = [Ag^+][Cl^-]$$

$$10^{-10} = [Ag^+]\,1$$

$$[Ag^+] = 10^{-10} \text{ M}$$

QUESTION – (AIIMS 2018 (E), 27 May)

The standard EMF for the cell reaction,
$Zn + Cu^{2+} \rightarrow Cu + Zn^{2+}$ is 1.1 volt at 25°C.
The EMF for the cell reaction, when 0.1 M Cu^{2+} and 0.1 M Zn^{2+} solutions are used, at 25°C is:

(A) 1.10 V

(B) 0.10 V

(C) –1.10 V

(D) –0.110 V

QUESTION (JEE Advance 2016)

For the following electrochemical cell at 298 K,

$$Pt_{(s)} \mid H_{2(g)}, (1\ bar) \mid H^+_{(aq)}, (1\ M) \parallel M^{4+}_{(aq)} \mid M^{2+}_{(aq)} \mid Pt_{(s)}$$

$E_{cell} = 0.092$ V when $\frac{[M^{2+}_{(aq)}]}{[M^{4+}_{(aq)}]} = 10^x$

Given: $E^o_{M^{4+}/M^{2+}} = 0.151$V, RT/F = 0.059 V. The value of x is:

A) −2

B) −1

C) 1

D) 2 ✓

$H_2 \rightarrow 2H^+ + 2e$

$M^{4+} + 2e \rightarrow M^{2+}$

$H_2(g) + M^{4+} \rightarrow 2H^+ + M^{2+}$

$E_{cell} = E^\circ_{cell} - \frac{0.059}{2} \log \frac{[H^+]^2 [M^{2+}]}{[M^{4+}] P_{H_2}}$

$0.092 = 0.151 - \frac{0.059}{2} \log \frac{[M^{2+}]}{[M^{4+}]}$

$+0.059 = +\frac{0.059}{2} \log \frac{[M^{2+}]}{[M^{4+}]}$

$\log \frac{[M^{2+}]}{[M^{4+}]} = 2$

$\frac{[M^{2+}]}{[M^{4+}]} = 10^2 = 10^x$

$x = 2$

QUESTION (JEE Advance 2017)

For the following cell, $Zn_{(s)} \mid ZnSO_{4(aq)} \parallel CuSO_{4(aq)} \mid Cu_{(s)}$ when the concentration of Zn^{2+} is 10 times the concentration of Cu^{2+}, the expression for ΔG (in J mol^{-1}) is [F is Faraday constant; R is gas constant; T is temperature; E°_{cell} = 1.1 V]

(A) 2.303 RT – 2.2F

(B) –2.2 F

(C) 2.303 RT + 1.1 F

(D) 1.1 F

$$\Delta G = \Delta G^\circ + 2.303 RT \log \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

$$= -2FX1.1 + 2.303 RT \log 10$$

$$= -2.2F + 2.303 RT$$

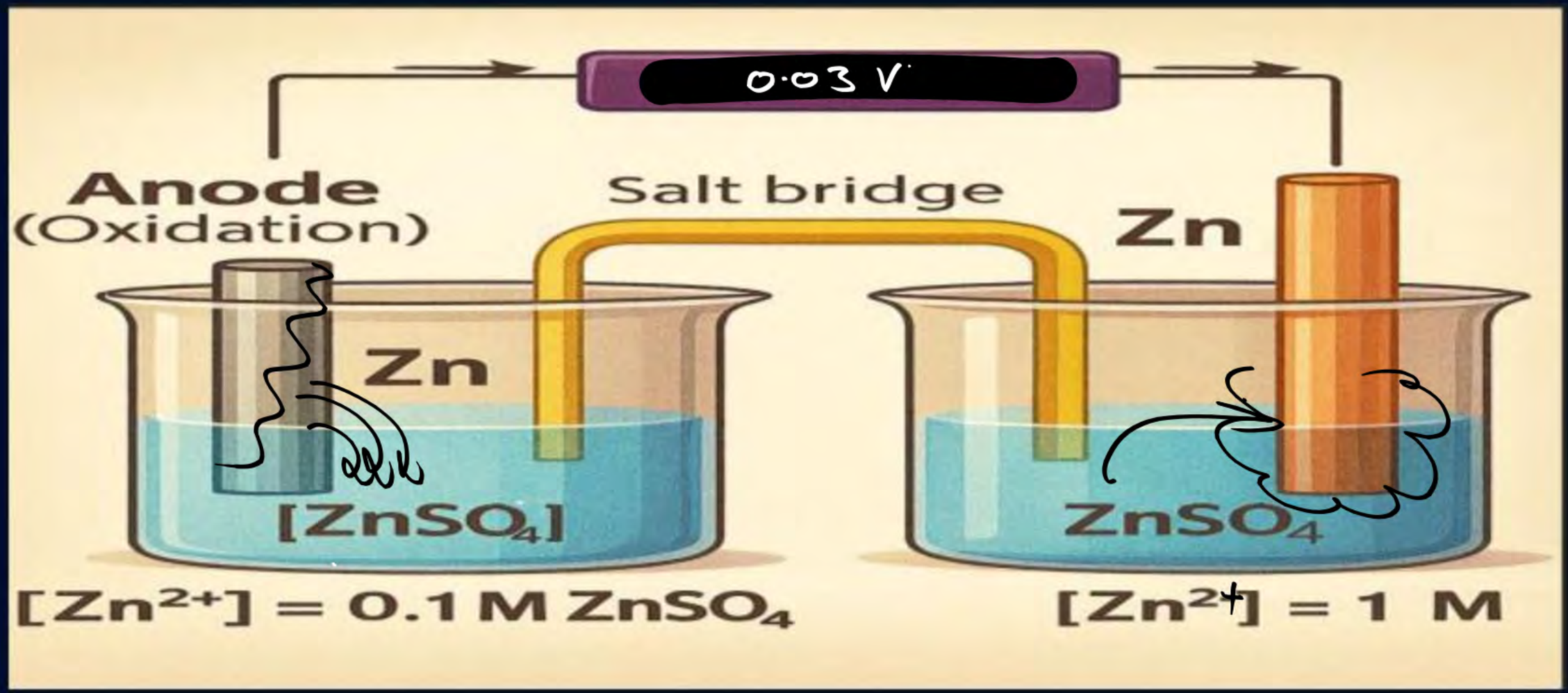
Concentration Cells

2 Types

① electrolyte Conc. Cells

② electrode Conc. Cells

Electrolyte Concentration Cells



MIT

① electrolyte Conc. Cell $\Rightarrow E^{\circ}_{cell} = 0$ V

② $E_{cell} = \frac{0.059}{n} \log \frac{[C]}{[A]}$

$\approx \frac{0.06}{n} \log \frac{[C]}{[A]}$

$[C]$ = Conc. Cathode elec

$[A]$ = —— Anode ,,

③ E_{cell} = (+)ve $\Rightarrow [C] > [A]$

$E^{\circ}_{cell} = E^{\circ}_{Zn^{2+}/Zn} - E^{\circ}_{Zn^{2+}/Zn}$

$E_{cell} = E^{\circ}_{cell} - \frac{0.059}{n} \log \frac{[A]}{[C]}$

$= \frac{0.059}{n} \log \frac{[C]}{[A]}$

QUESTION

0.009

Find E_{cell} of following:

$Zn(s) \mid Zn^{2+}(0.1\ M) \mid\mid Zn^{2+}(0.2\ M) \mid Zn(s)$

$$E_{cell} = \frac{0.06}{2} \log \frac{0.2}{0.1}$$

$$= 0.03 \times 0.3$$

$$= 0.009\ V$$

MIT

electrolyte Conc Cell for Hydrogen electrode

$$E_{cell} = \frac{0.06}{1} \log \frac{[H^+]_C}{[H^+]_A}$$

$$= 0.06 \left[P_H \text{ anode} - P_H \text{ Cathode} \right]$$

$$E_{cell} = \frac{0.06}{n} \log \frac{[H^+]_C}{[H^+]_A}$$

$$= 0.06 \left[\log [H^+]_C - \log [H^+]_A \right]$$

$$= 0.06 \left[-pH_{Cat} + pH_{Anod} \right]$$

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

QUESTION

Find E_{cell} of following:

$Pt, H_2 \mid pH = 5 \parallel pH = 3) \mid H_2, Pt$

Ans $E_{cell} = 0.12\ V = 0.118\ V$

$$E_{cell} = \frac{0.06}{n}\left(pH_{anode} - pH_{cathode}\right)$$

$$= \frac{0.06}{1}\left(5 - 3\right)$$

$$= 0.06 \times 2 = 0.12\ V$$

Electrolyte Concentration Cells

$Pt, H_2(g)_{P_1} / H^+ / H_{2(P_2)}(g)$

$n = 2$

anode

$H_2(g) \rightarrow 2H^+ + 2e$

(P_1) $\quad$ 0.1

Cathode

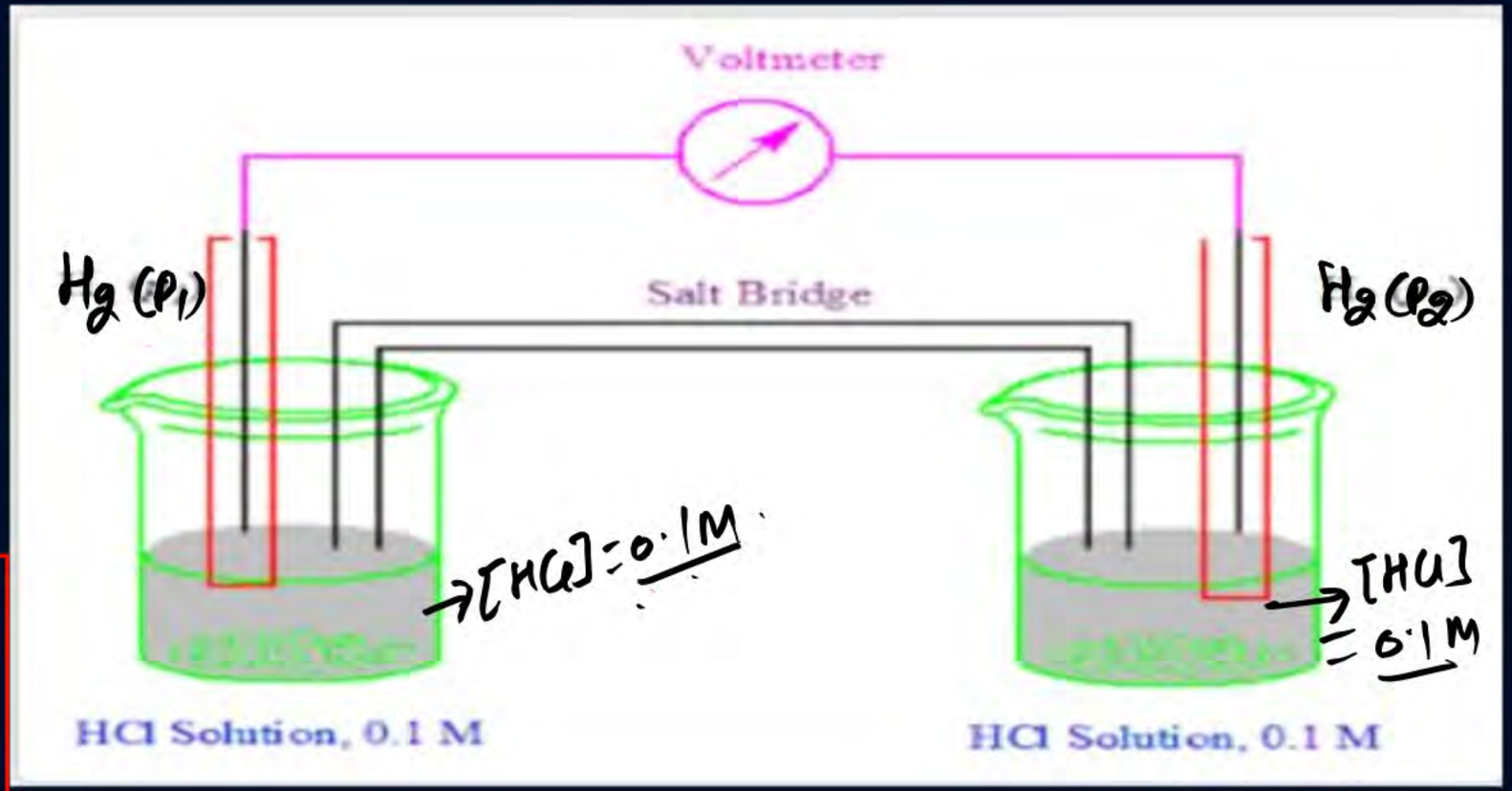
$2H^+ + 2e \rightarrow H_2(g)$

$(0.1M) \quad P_2$

$H_2(g) \rightarrow H_2(g)$

$P_1 \quad P_2$

$$E_{cell} = + \frac{0.06}{2} \log \frac{(P_{H_2})_A}{(P_{H_2})_C}$$



QUESTION

Calculate E_{cell}

Pt, H_2 (5 atm)/ H^+/ H_2 (2 atm), Pt

AB

$$E_{cell} = 0.03 \log \frac{5}{2}$$

$$= 0.03 (\log 5 - \log 2)$$

$$= 0.03 (0.7 - 0.3)$$

$$= 0.03 \times 0.4$$

$$= 0.012\ V$$

When Electrode pressure and Electrolyte conc. Both are different:

for $H_2(g)$

$$E_{cell} = \frac{0.06}{2} \log \frac{(P_{H_2})_A [H^+]_C^2}{(P_{H_2})_C [H^+]_A^2}$$

anode : $H_2(g) \rightarrow 2H^+ + 2e^-$

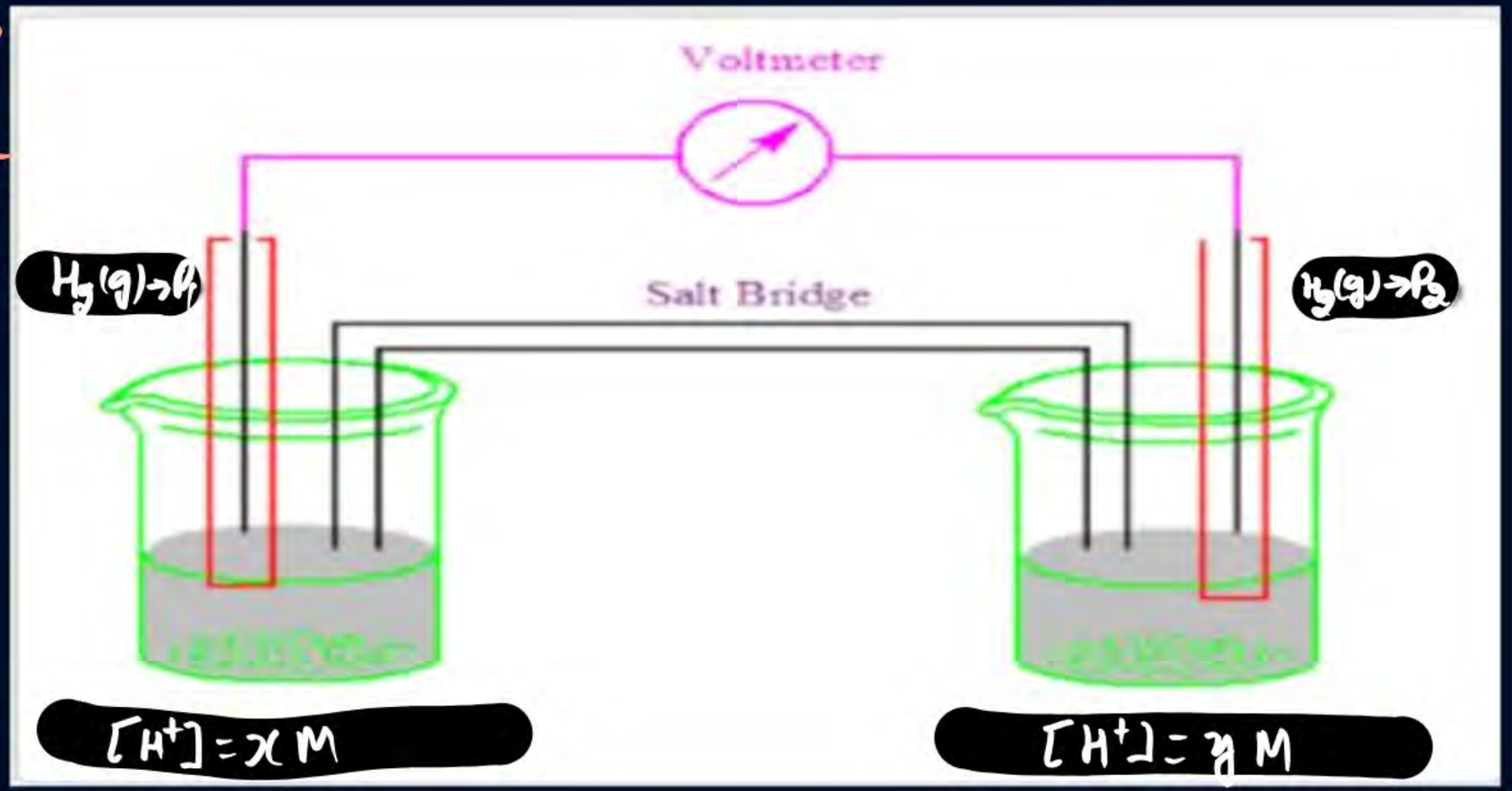
P_1 $\quad$ x M.

Cathode $2H^+ + 2e^- \rightarrow H_2(g)$

yM $\quad$ P_2

$$H_2(g) + 2H^+ \rightarrow 2H^+ + H_2(g)$$

P_1 $\quad$ yM $\quad$ xM $\quad$ P_2



QUESTION

Find E_{cell} of the following:

Pt, H_2 (5 atm) | H^+ (1 M) || H^+ (2 M) | H_2 (25 atm), Pt

$n = 2$

$$E_{cell} = 0.03 \log \frac{5 \times (2)^2}{\frac{25}{5} \times (1)^2}$$

$$= 0.03 [\log 4 - \log 5]$$

$$= 0.03 [0.6 - 0.7]$$

$$= -0.003 V$$

How Cells are Discharged ?

$E^{0}_{Zn^{2+}/Zn} = -0.76$

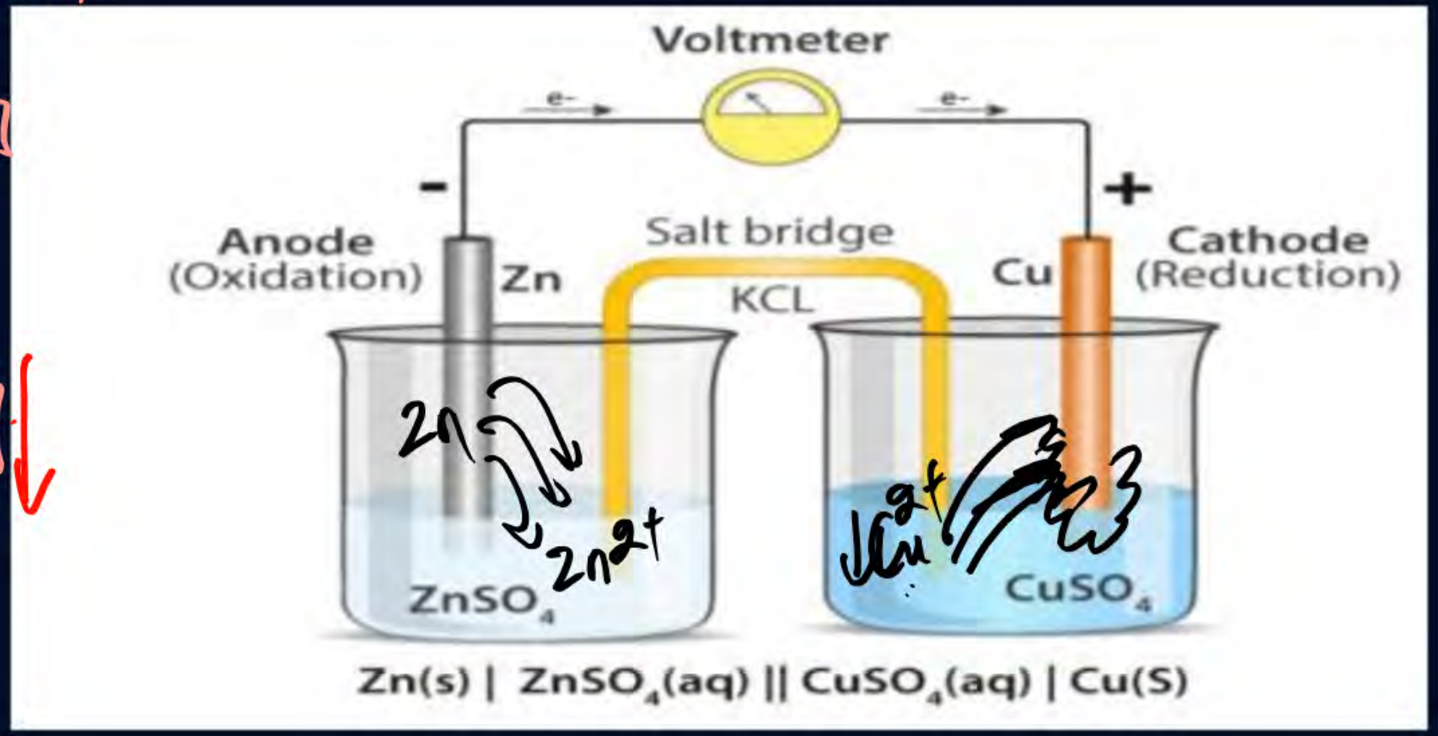
$E^{0}_{Zn/Zn^{2+}} = 0.76$

$E_{Zn/Zn^{2+}} = E^{0}_{Zn/Zn^{2+}} - \frac{0.06}{2} \log [Zn^{2+}]$

$= 0.76 - 0.03 \log [Zn^{2+}]$

$E_{Cu^{2+}/Cu} = E^{0}_{Cu^{2+}/Cu} + \frac{0.06}{2} \log [Cu^{2+}]$

$= 0.34$



#MIT

Cells discharge.

$$E_{Zn/Zn^{2+}} = E_{Cu^{2+}/Cu}$$

$$E_{Cell} = 0$$

QUESTION

$0.059 \approx 0.06$

$\frac{111}{3} = 37$

The cell, $Zn \mid Zn^{2+}$ (1 M) || Cu^{2+} (1 M) | Cu (E°_{cell} = 1.10 V), was allowed to be completely discharged at 298 K. The relative concentration of Zn^{2+} to Cu^{2+} $\left(\frac{[Zn^{2+}]}{[Cu2+]}\right)$ is:

A. Anti log (24.08)

B. antilog (37.3)

C. 1037.3

D. 9.65×10^4

$$E_{cell} = E^{\circ}_{cell} - \frac{0.06}{2} \log \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

$$+1.1 = +0.03 \log \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

$$\frac{1.10}{0.03} = \log \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

$$\text{antilog}(37) = \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

Question

For the following cell at 298 K

Ag | AgI (saturated) | AgI(s) | Ag;

Calculate solubility product of AgI at 298 K.

$\frac{2.303RT}{F} = 0.06V$

K_{sp} = eq. Constt.

eq. $\Rightarrow \Delta G = 0 = -nFE_{cell}$

Anode :- $Ag \rightarrow Ag^{+} + 1e^{-}$

Cathode :- $AgI + 1e \rightarrow Ag + I^{-}$

$AgI \rightarrow Ag^{+} + I^{-}$

$n = 1$

$E^{\circ}_{cell} = -0.960\ V$

$K_{sp} = [Ag^{+}][I^{-}]$

$E_{cell} = E^{\circ}_{cell} - \frac{0.06}{n}\log [Ag^{+}][I^{-}]$

$0 = -0.96 - \frac{0.06}{1}\log K_{sp}$

$-\frac{0.96}{0.06} = \log K_{sp}$

$-16 = \log K_{sp}$

$K_{sp} = \text{antilog}(-16)$

$K_{sp} = 10^{-16}$

Home work from modules

Parikshit ⇒ Q1, 3, 6, 10, 11, 12

after Complete revision of electrochemistry attempt these

THANK YOU