

VICTORY ZONE CHEMISTRY CLASSES

ADDRESS - 1ST FLOOR PRADEEP PLAZA, IN FRONT OF CHOPRA PALACE, ADARSH NAGAR DURG

MO 9424168056,8269055301

CHAPTER SCHEDULE

ELECTROCHEMISTRY

3 JUNE 2025 TO 6 JULY 2025

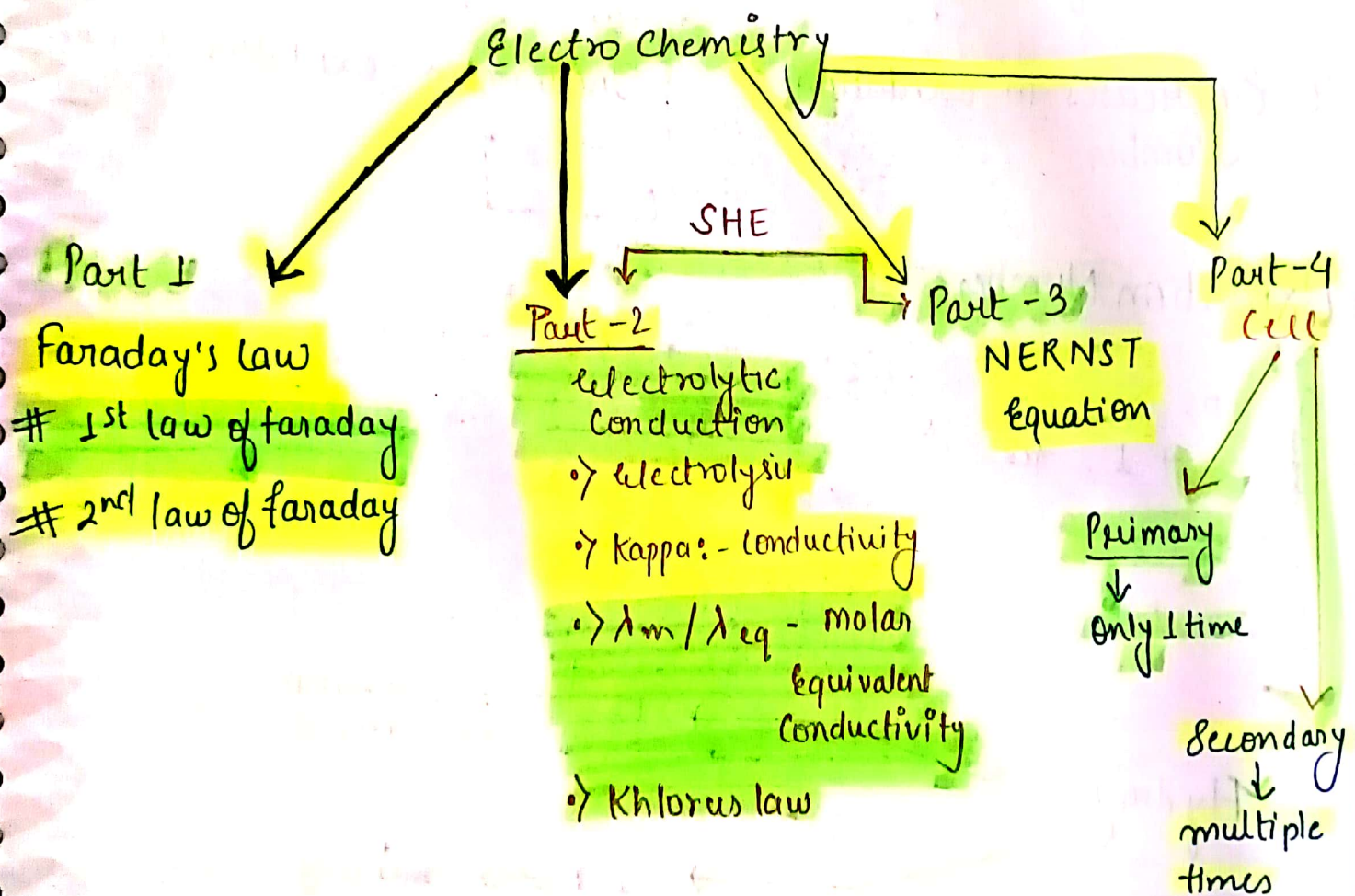
S. NO	CLASS DETAIL	TOPICS DETAILS	CHECK LIST
1.	LECTURE NO. 1	REDOX REACTION 30Q	
2.	LECTURE NO. 2	DANILE CELL	
3.	LECTURE NO. 3	DANILE CELL & ELECTROLYTIC CELL DIFFERENCE	
4.	LECTURE NO. 4	SHE 20Q	
5.	LECTURE NO. 5	REDUCTION POTENTIAL 50Q	
6.	LECTURE NO. 6	NERNST-EQUATION 30Q	
7.	LECTURE NO. 7	EQUILIBRIUM, THERMO & ELECTRO RELATION	
8.	LECTURE NO. 8	EQUILIBRIUM, THERMO & ELECTRO RELATION 20Q	
9.	LECTURE NO. 9	ELECTROLYTIC CONDUCTION THEORY	
10.	LECTURE NO. 10	ELECTROLYTIC CONDUCTION 30Q	
11.	LECTURE NO. 11	KOHLROUSH LAW	
12.	LECTURE NO. 12	KOHLROUSH LAW 20Q	
13.	LECTURE NO. 13	DEGREE OF IONIZATION 15Q	
14.	LECTURE NO. 14	FARADY'S 1 ST & 2 ND LAW THEORY	
15.	LECTURE NO. 15	FARADY'S 1 ST & 2 ND LAW 30Q	
16.	LECTURE NO. 16	BOARDS PYQS DISCUSSION	
17.	LECTURE NO. 17	COMPLETE CHAPTER QUESTIONS PYQS	
18.	LECTURE NO. 18	SELF ANALYSIS PAPER IN CLASS	
19.	TEST NO. 5	05/07/2025 MCQS 30Q MM:120 (YOUR MARKS)	
20.	TEST NO. 6	06/07/2025 THEORY 16Q MM:30 (YOUR MARKS)	
21.	HOME WORK	NUMBER OF QUESTION DONE BY YOUR SELF	
22.	HOME WORK	NCERT BOOK READING & BACK QUESTIONS	

JOIN 'VICTORY ZONE' FOR BEST CONTENT AND CONCEPT



CHAPTER = 2

ELECTRO CHEMISTRY



* Redox Reaction

Reduction	Oxidation
1) Addition of Hydrogen	Addition of Oxygen
2) Removal of Oxygen	Removal of Hydrogen
3) Gain of electron $[e^-]$.	Loss of electron $[e^-]$
4) Decreases in Oxidation Number	Increase in Oxidation Number.

Oxidation Number

Groups

1:- +1, ~~+2~~

2:- +2

13:- +3

Hydrogen

+1

→ Non metal
Ex - HCl

-1

→ metal
Ex - NaH

*
Non-metal

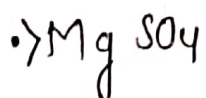
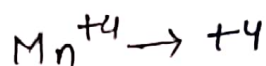
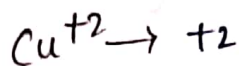
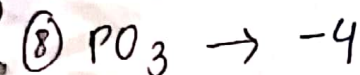
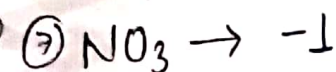
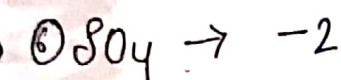
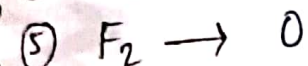
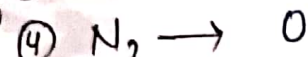
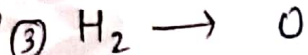
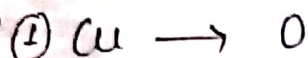
Solve the Linear Equation

$$1 + 2x + 3$$

$$x = -2$$

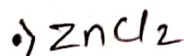
Example

* Self Bonding Oxidation Number = 0



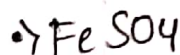
$$+1 + x - 2(4) = 0$$

$$x = +7$$



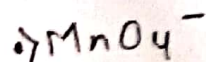
$$x - 1(2)$$

$$x = +2$$



$$x - 2 = 0$$

$$x = 2$$



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

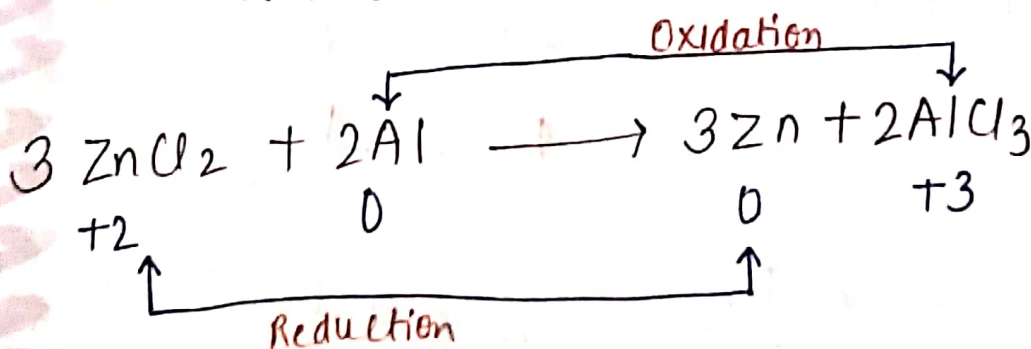
$$x = +7$$

* Redox Reaction

Increase in Oxidation No.



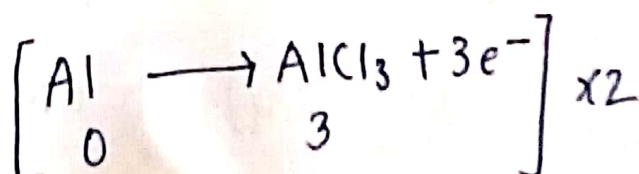
Decrease in Oxidation Number



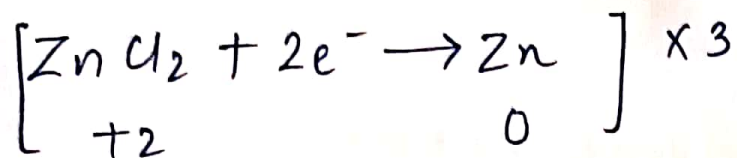
Electron Balancing Method

Half Reaction

Oxidation Half Reaction

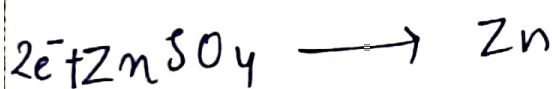
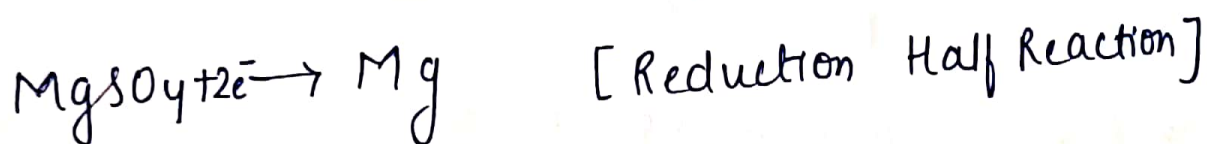
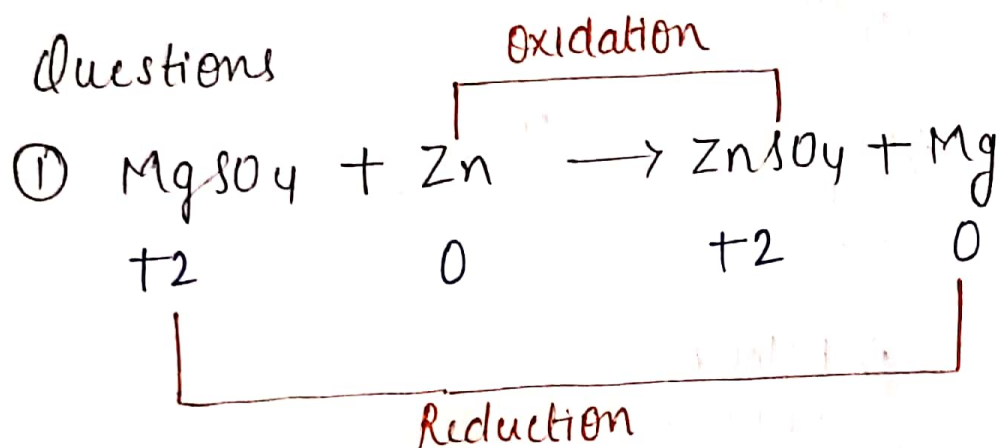


Reduction Half Reaction

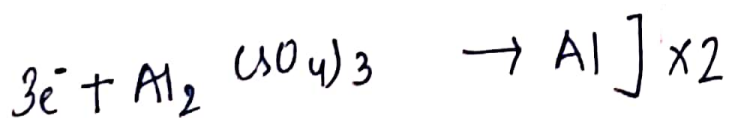
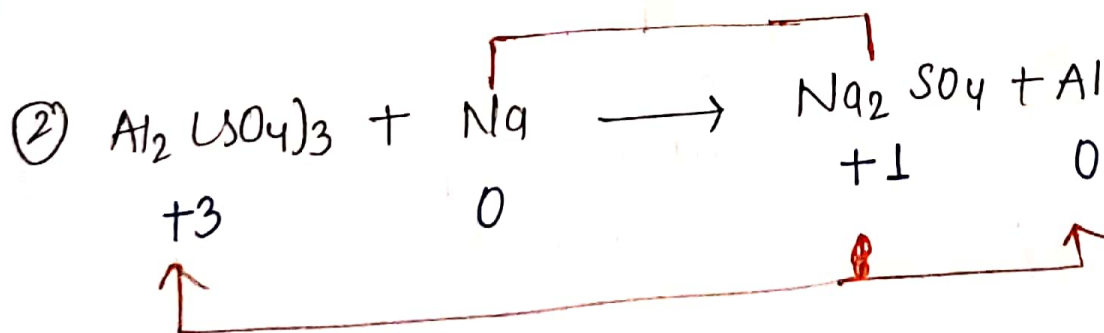


So Total no. of $e^- = 6$

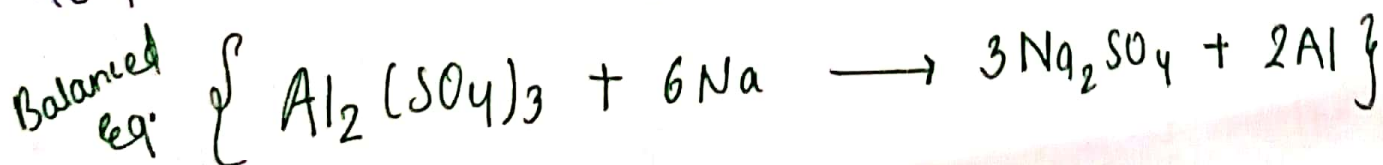
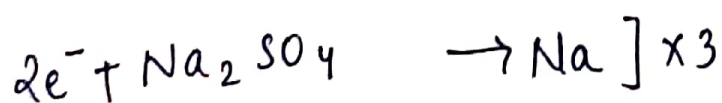
Questions

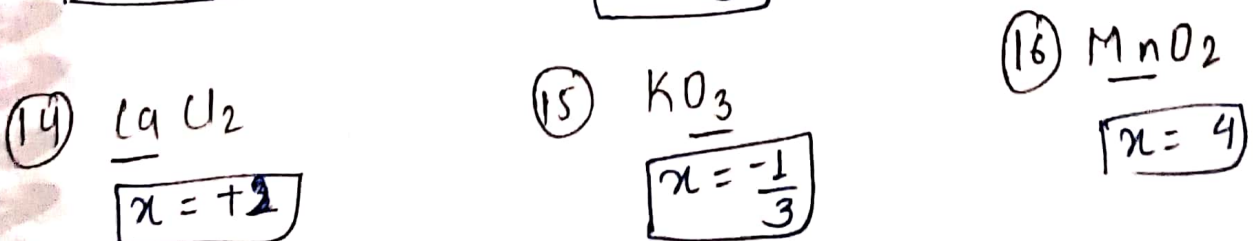
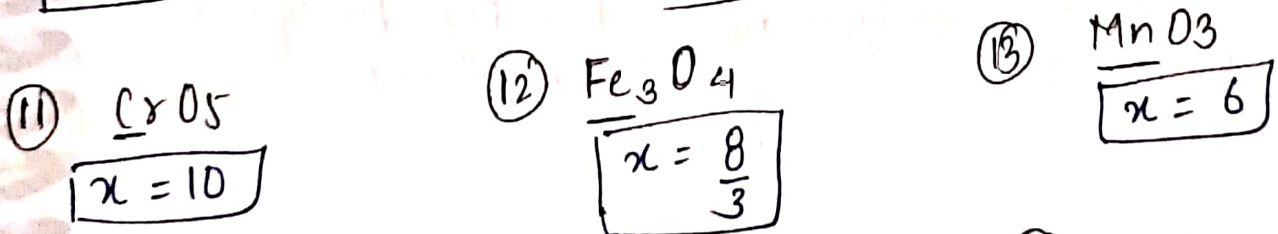
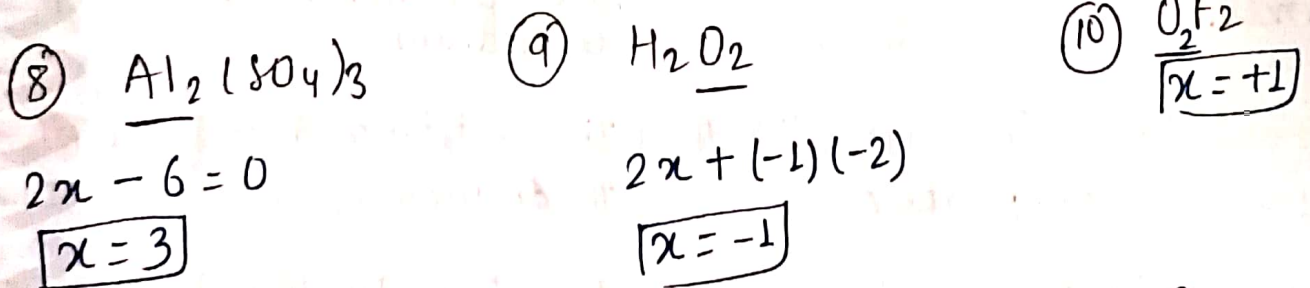
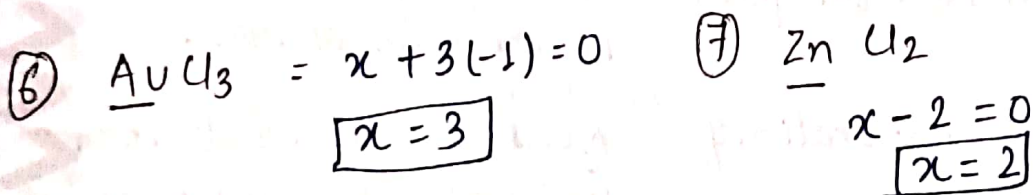
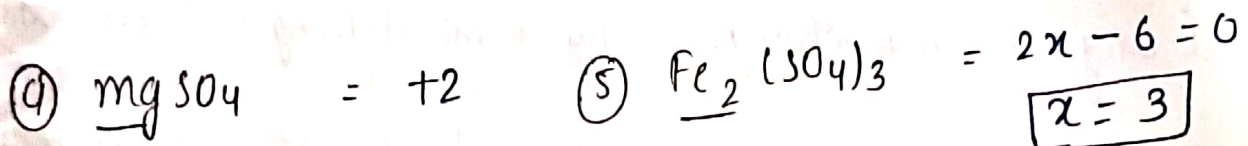
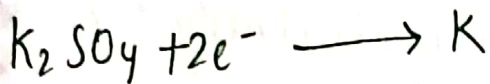
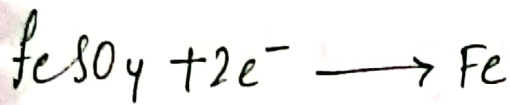
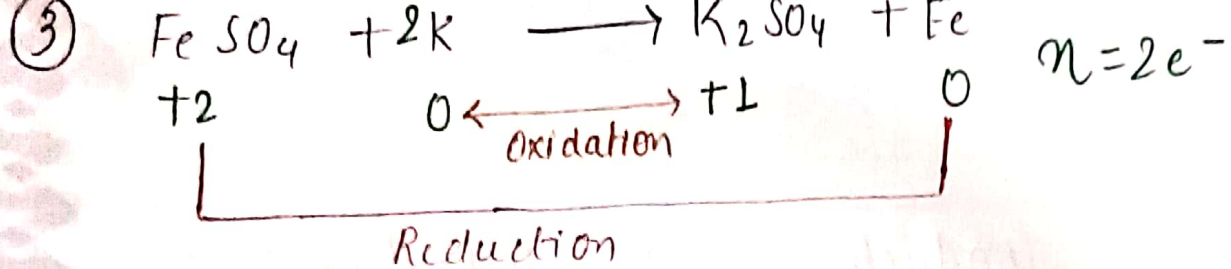


So, Total no. of electron = $2e^-$



no. of electron = $6e^-$

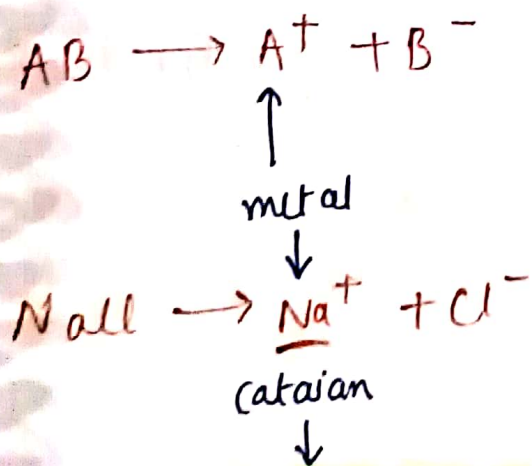
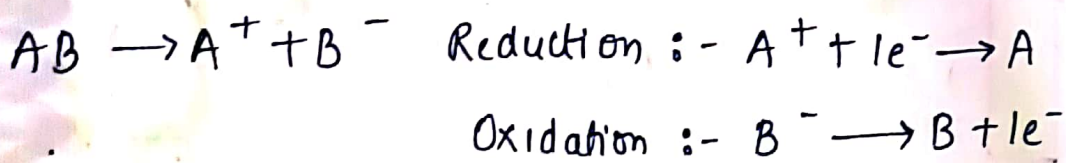
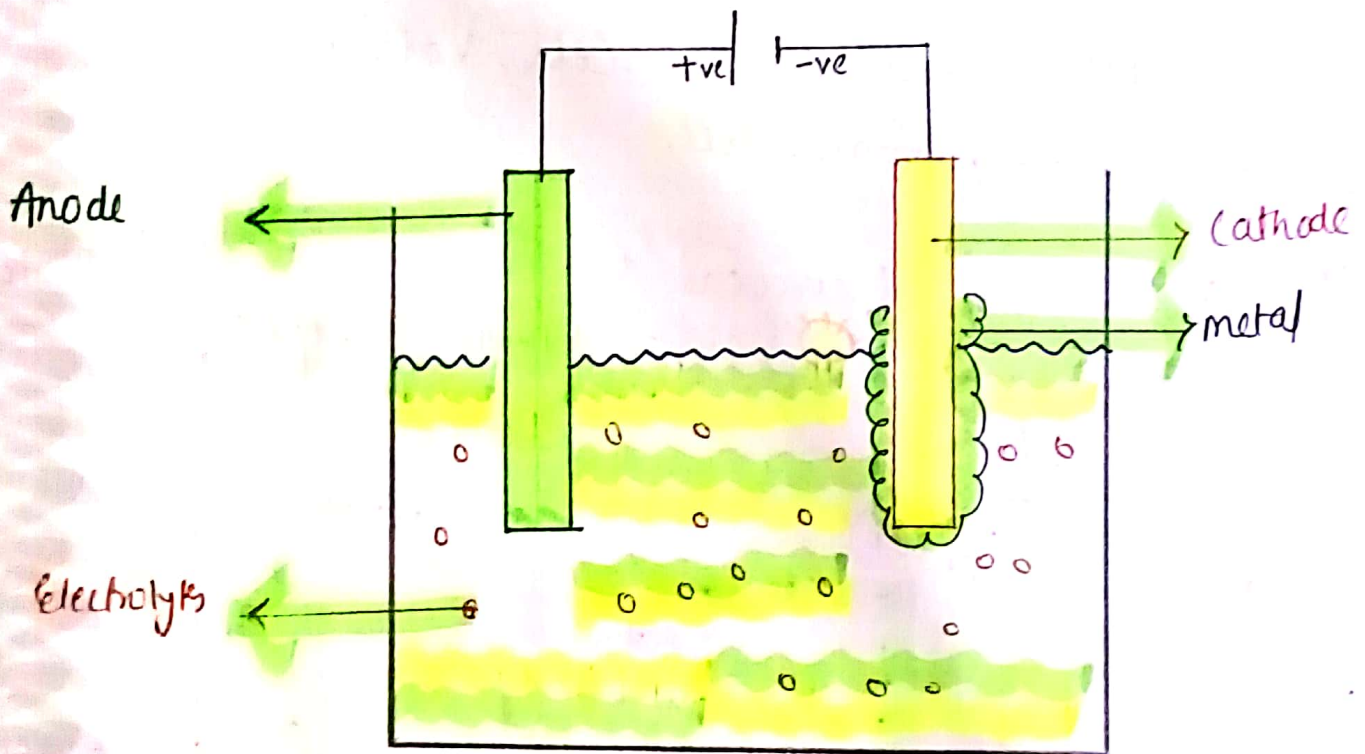




Difference

Metallic Conductor	Electrolytic Conductor
Charge carriers are free Electron. Flow of electricity takes place without any chemical change. No transfer of matter take place Resistances is due to vibration of kernels. The resistance increases with the increase in temperature Faraday's law of Electrolysis are not followed	Charge carriers are free ions Flow of Electricity takes places by chemical changes at electrodes Transfer of matter takes places in the form of ions. Resistance is due to inter ionic attraction and viscosity of medium The resistance decreases with the the increase in temperature Faraday's law of electrolysis are allowed

Chemical energy $\rightleftharpoons$ Electrical energy
electrolysis



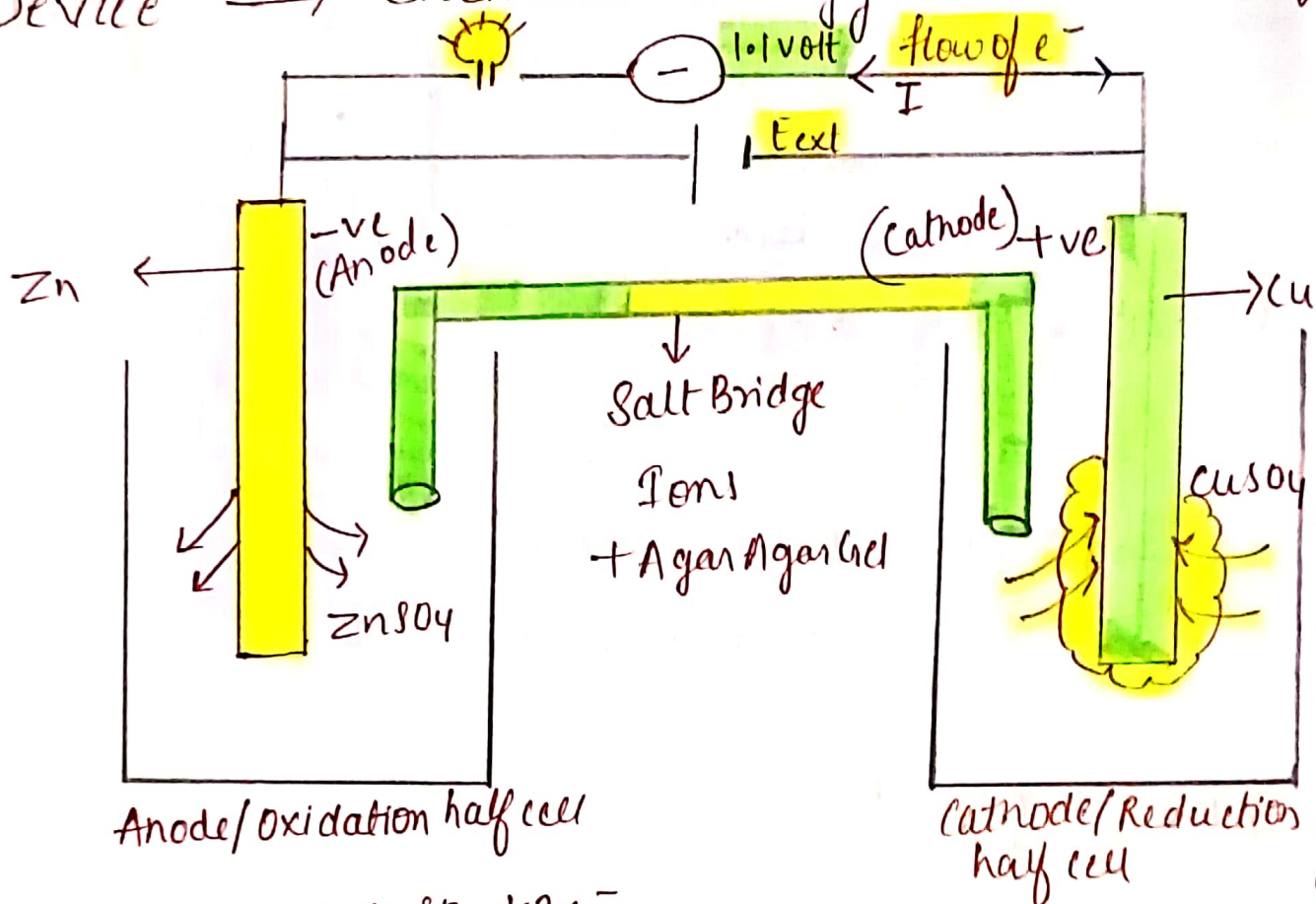
Cation की तरफ attract होने वाला ion

5/5/2025

ELECTROCHEMICAL CELL

It is also called Volatic Cell, Voltaic cell, Galvanic cell, Daniell cell.

Device $\rightarrow$ Chemical Energy $\rightarrow$ Electrical Energy



A L O N E

A :- Anode

L :- Left

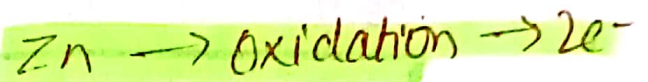
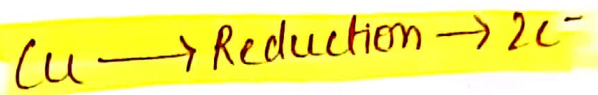
O :- Oxidation

N :- Negative

Functions

① Salt Bridge

- Electrical Neutral
- Circuit Complete



→ Cu → Deposited [thick]

→ Zn → Dissolved [thin]

flow of e^- :- Zn → Cu [-ve → +ve]

flow of I :- Cu → Zn [+ve → -ve]

Cases :- I) When you Remove Salt Bridge

→ Circuit Bridge Break It will not work

→ No. of e^- flow will stop

Cases - II) Dead After some times

CuSO_4 (blue) → Transparent

↳ All Cu will deposited

∴ Reason No more Cu Available for deposited

Cases - III) When Electric external = 1.1 Volt

No. electron will flow

Equilibrium State

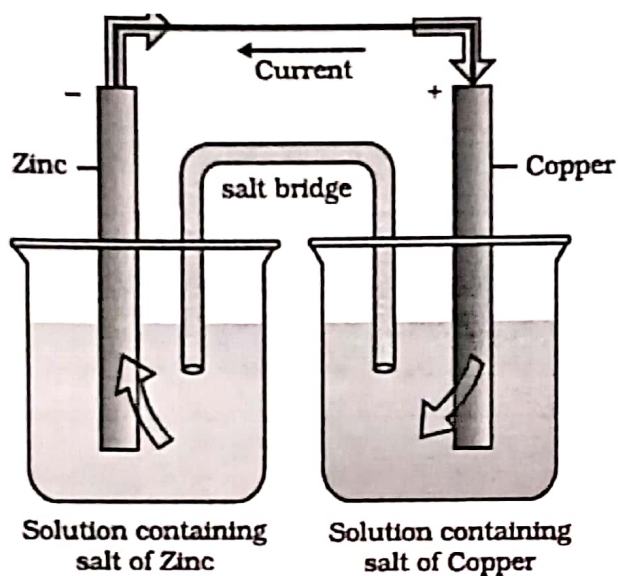
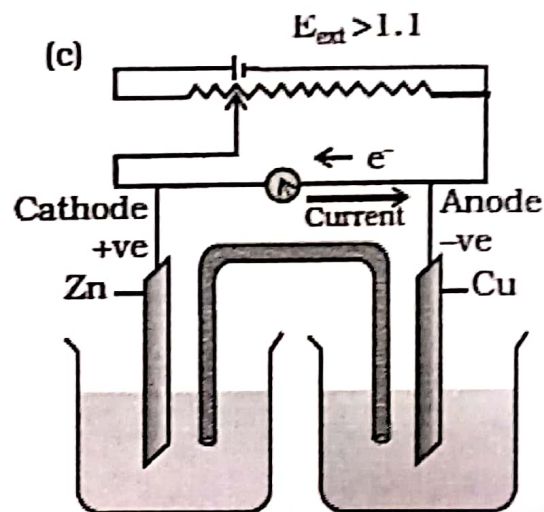


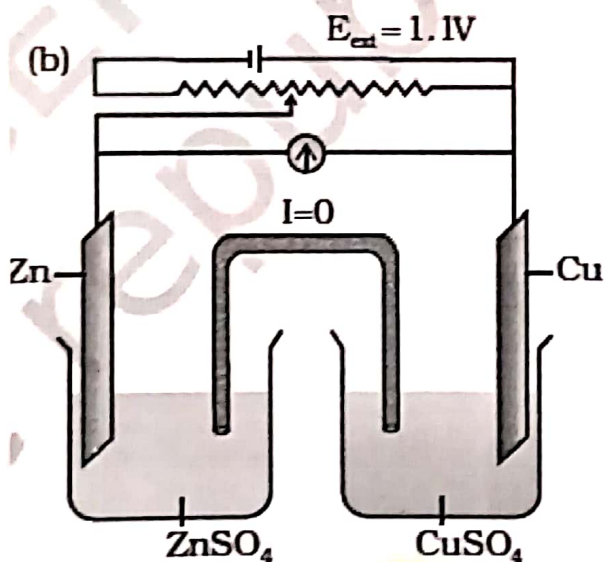
Fig. 2.1: Daniell cell having electrodes of zinc and copper dipping in the solutions of their respective salts.

Case - V Charging

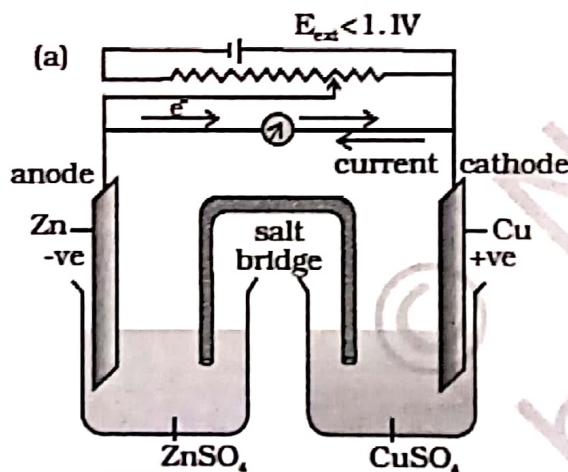


- Flow of e^- $Cu \rightarrow Zn$
- Flow of I $Zn \rightarrow Cu$
- Zn is deposited at anode
- Cu is dissolved at cathode

Case - III



- 0. Current Flow
- 0. Electron Flow
- 0. Chemical Reaction



Case IV

- Flow of e^- $Zn \rightarrow Cu$
- Flow of I Cu to Zn
- Cu deposited at cathode
- Zn ~~deposited~~ dissolved at anode

* Cell Representation

IUPAC

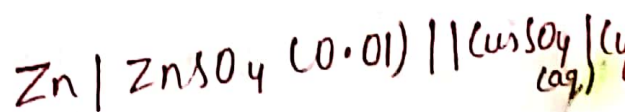
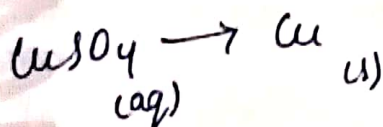
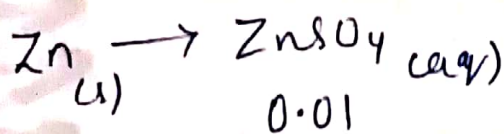
→ We Required to half cell to produce an electrochemical cell which can be represented by following few Rules given by IUPAC:-

- ① The Anode Half cell is always written on left and Cathode half cell written by Right side.
- ② The Separation of two phases is shown by Vertical line.
- ③ The Salt Bridge is represented by dash $||$ Double Vertical line.
- ④ The Significant feature of Substance like pressure of gas, concentration of ion etc. are indicated in the bracket immediately [atq] After writing the substance.

Anode Salt Bridge Cathode

Anode || Cathode

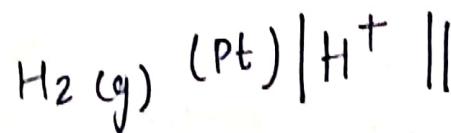
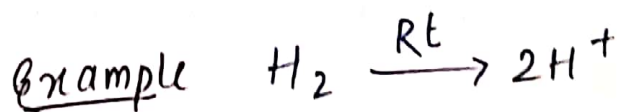
Daniel Cell



→ Catalyst ()

→ Convinient () / (concentration [])

→ State (solid, liq, gas)



Salt Bridge and Its Functions

- ① Salt Bridge is inverted U-Tube containing a solution of Inert electrolyte with Agar-Agar gel and Gelatine to converted into semi solid.
- ② Ions of Inert electrolyte do not involve in any chemical change
- ③ The electrolyte in salt bridge have a equal speed of cation and anion.
- ④ Example :- KCl , KNO_3 , NH_4NO_3 etc
- ⑤ If Ag^+ , Hg^{+2} , Pb^{+2} , Tl^+ ions are present then KCl is not used. Because they form precipitation with ions. Example :- $AgCl$, $HgCl_2$, $PbCl_2$, $TlCl$ etc.

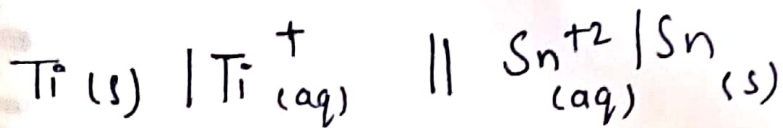
Functions

- It connect the solution of two half cell to complete the circuit.
- It maintain the electrically neutral of the solution in order to give continuous flow or generation of current
- If salt bridge removed the voltage dropped to zero.
- It also prevent salt-salt junction.

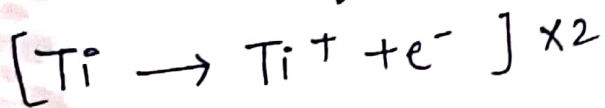
Salt Salt Junction:- The Potential Difference which arise between two solution where cation and anion are direct contact to each other.

(conversion)

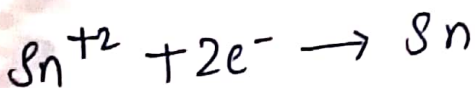
Over Rxn $\rightleftharpoons$ cell IUPAC representation



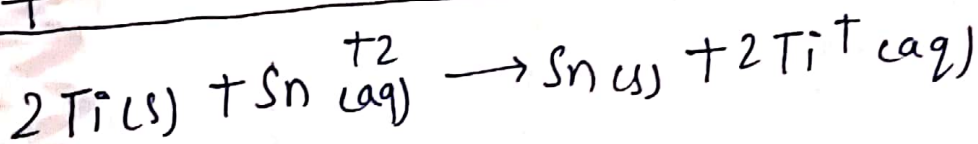
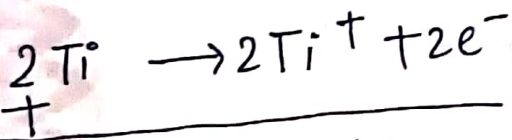
Oxidation Half Reaction



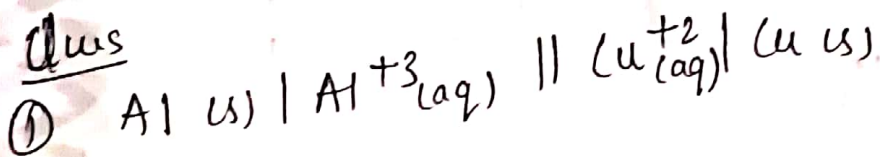
Reduction half Reaction



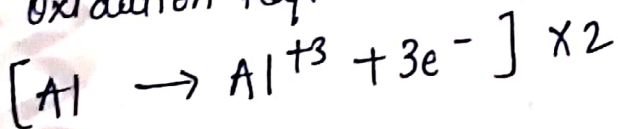
$$\text{no. of } e^- = 2 = n$$



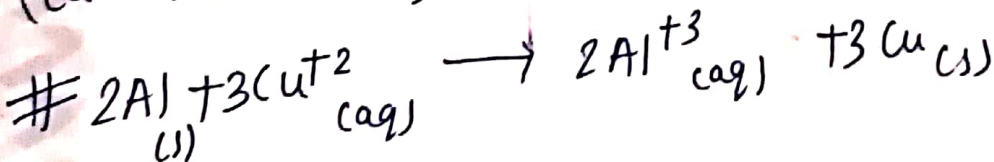
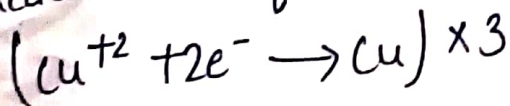
Ques



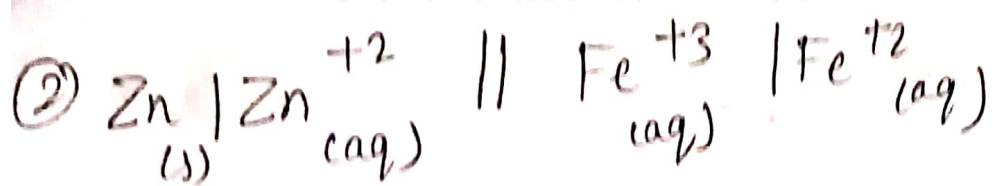
Oxidation Half Reaction



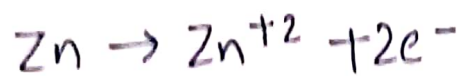
Reduction Half Reaction



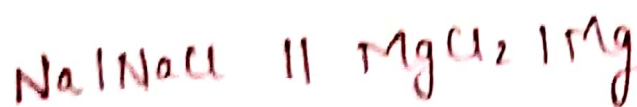
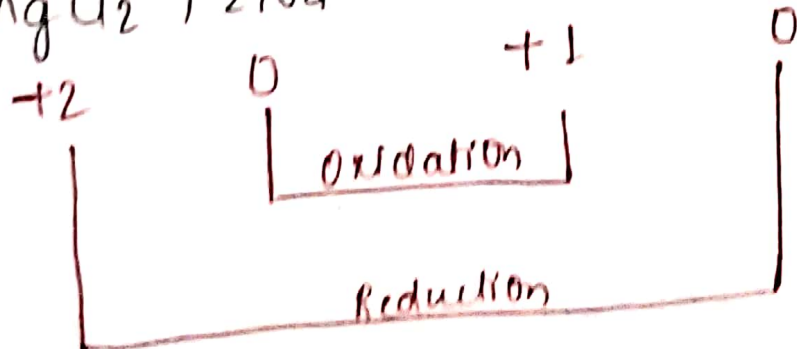
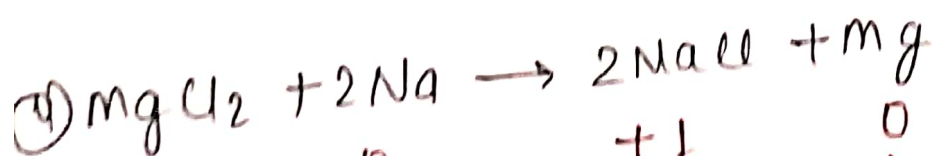
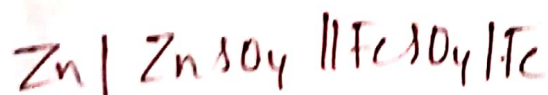
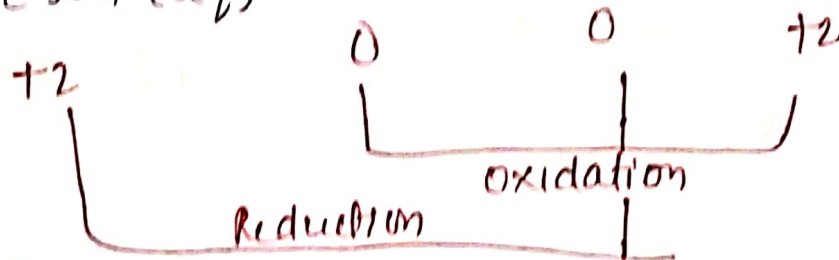
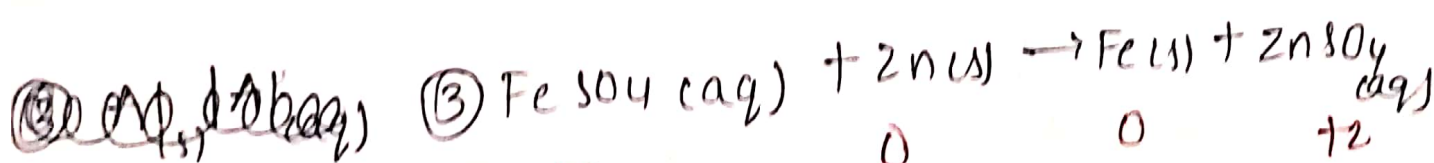
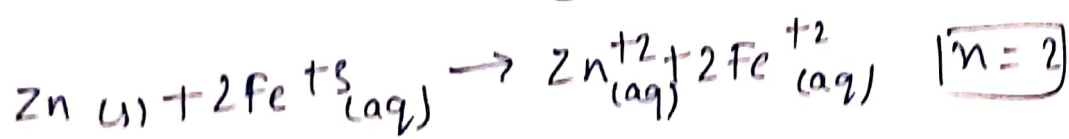
$$\boxed{n = 6}$$



Oxidation Half Reaction



Reduction Half Reaction



* Most Imp 100 Stars

Electrochemical cells extensively used for determining the pH of solution, solubility product, equilibrium constant and other thermodynamics properties and for potentiometric titrations.

STANDARD HYDROGEN ELECTRODE POTENTIAL

SHE stands for Standard Hydrogen Electrode Potential

By convention its electrode potential is taken as zero.

Construction:-

① SHE consists of a Platinum Electrode coated with platinum black. The Electrode is dipped in an Acidic Solution and pure Hydrogen is bubbled through it.

② The Concentration of both reduced and oxidised

③ Forms of Hydrogen is maintained at unity.

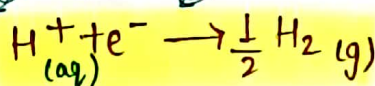
④ Example - (i.e) pressure of H_2 gas is 1 bar and concentration of Hydrogen ions is 1 molar.

Working

The Reaction take place in SHE is

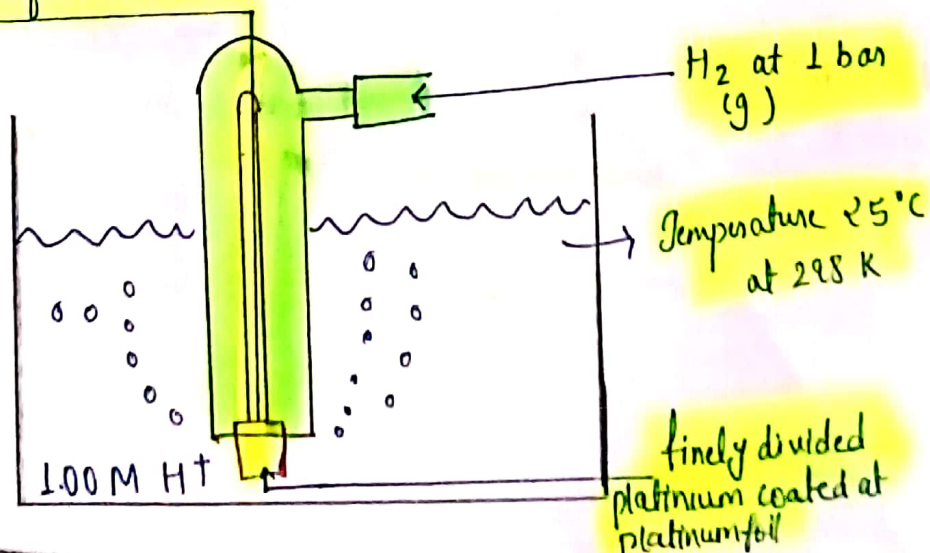
~~2H⁺ + 2e⁻ → H₂~~

~~2H⁺ + 2e⁻ → H₂~~ at 298 Kelvin



It is represented as $Pt(s) | H_2(g) | H^+(aq)$

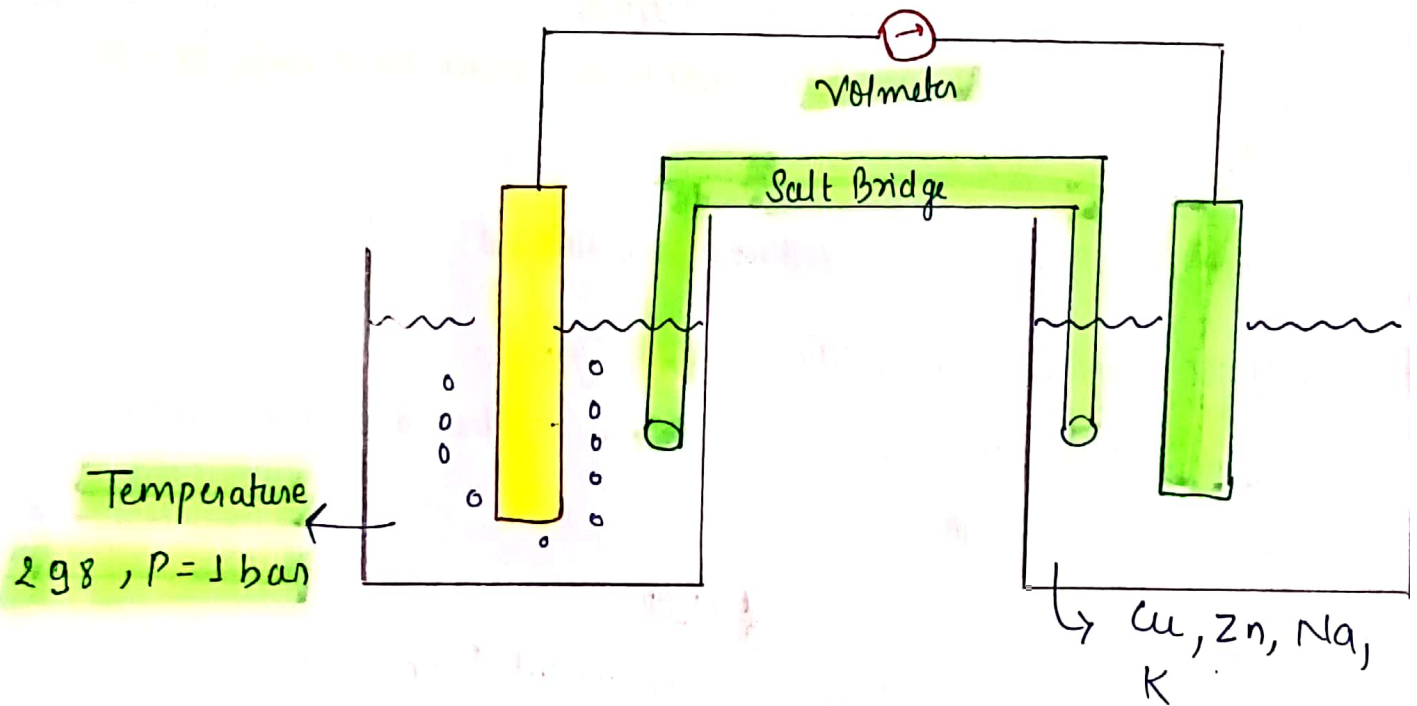
At 298 Kelvin the EMF of cell construction by taking SHE as a reference to zero



* N H E

→ When SHE as a reference point taken as zero SHE is called Normal Hydrogen Electrode.

* Reduction Potential



* Electrochemical Series

Like Kings Can Not Make A zebra cage In
 Li K Ca Na Mg Al Zn Cr Iron Fe²⁺
 Cold Night To Pop Hydrogen Cuckoo. I
 Cd Ni (Sn) (Pb) (H) (Cu) (I)
 Fainted As He Brought Chlorinated Gold fore me.
 (Fe) (Ag) (Hg) (Br) (Cl) (Au) (F)

Cu = +0.337 V
 Zn = -0.763 V
 Na = -2.71 V
 K = -2.925 V

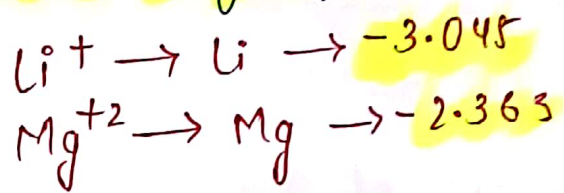
Application of Reduction potential or Electro Chemical Series

→ Top of the series are good reducing agent.

See
R → Right

R → Reducing Agent

-ve → right are best



Li is the best reducing agent.

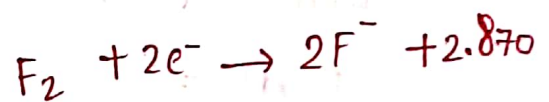
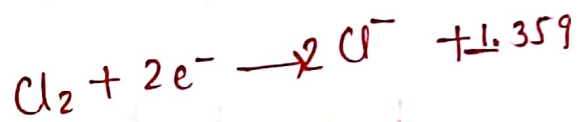
→ Down of the series are good Oxidizing Agent

see

L → left

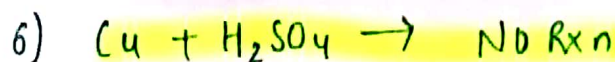
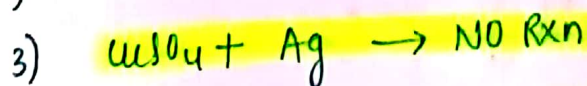
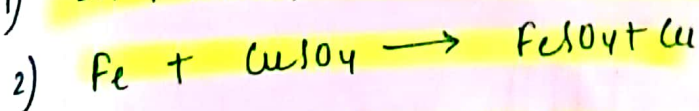
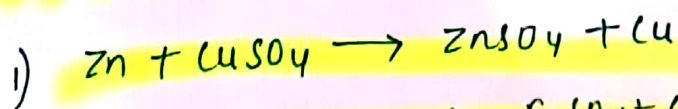
O → Oxidation

Value → must +ve are
Oxidised Agent



→ ऊपर वाले नीचे वाले को Replace करते हैं।

feasible :-



→ Top of series are electronegative

→ Down of the series are electronegative.

EMF of a cell

Electromotive force it is also called cell potential. It is repeated as E_{cell} , E_{mf} . It is the difference between the two electrode potential.

$$E_{\text{cell}} = E_{\text{cat}} - E_{\text{an}}$$

$$E_{\text{cell}} = E_{\text{red}} - E_{\text{ox}}$$

$$E_{\text{cell}} = E_{\text{right}} - E_{\text{left}}$$

$E^{\circ}_{\text{cell}} = -ve \rightarrow$ Non-feasible Reaction

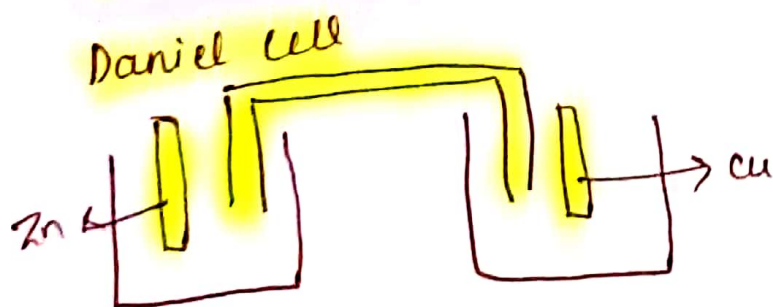
$\begin{matrix} \text{Cat} \rightarrow \text{An} \\ \text{An} \rightarrow \text{Cat} \end{matrix} \quad \left. \vphantom{\begin{matrix} \text{Cat} \rightarrow \text{An} \\ \text{An} \rightarrow \text{Cat} \end{matrix}} \right\} \text{Replacement of electrodes}$

EMF of Daniell cell

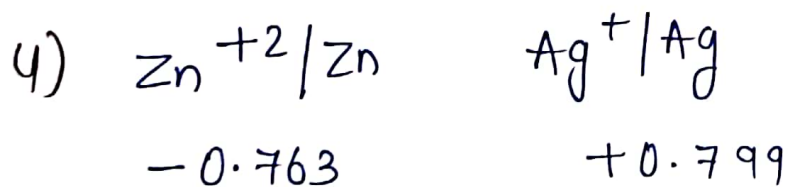
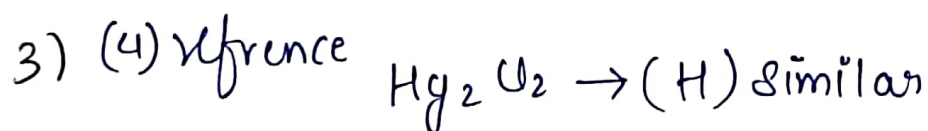
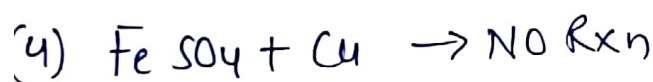
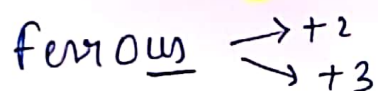
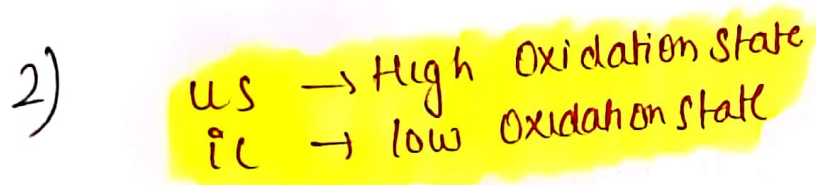
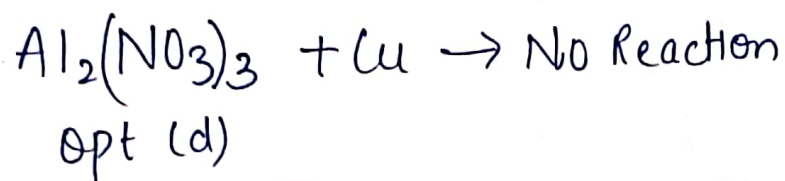
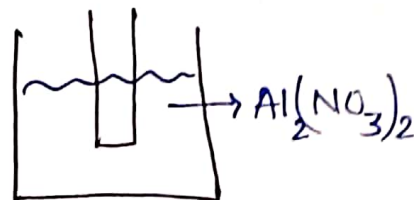
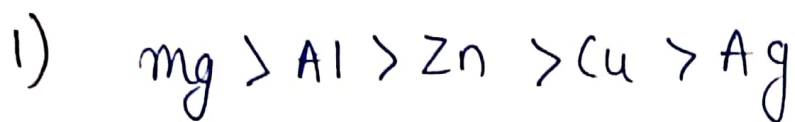
$$E^{\circ}_{\text{cell}} = E_{\text{Cu}^{2+}|\text{Cu}} - E_{\text{Zn}^{2+}|\text{Zn}}$$

$$E^{\circ}_{\text{cell}} = 0.337 - (-0.763)$$

$$E^{\circ}_{\text{cell}} = 1.10 \text{ volt}$$



DPP-04



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

$E^\circ_{cell} = E_{cat} - E_{an}$ $\hookrightarrow$ -ve less than E_{cat}
 $\downarrow$

+ve
more than E_{an}

$E^\circ_{cell} = 0.799 - (-0.763)$

$E^\circ_{cell} = 0.799 + 0.763$

$E^\circ_{cell} = 1.56 \text{ volt}$

5) Red $\rightarrow$ more -ve
 $\hookrightarrow$ Right side

So, Zn opt (1)

6) (2) Zn < Cu
-ve +ve

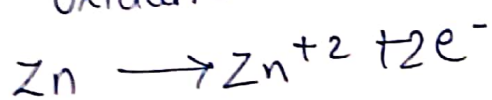
7) (4) Ag

8) (3) D

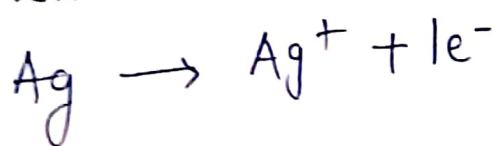
10/6/25

9) Two question

Oxidation Potential

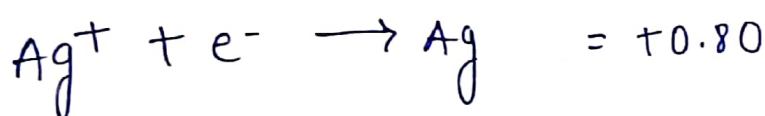
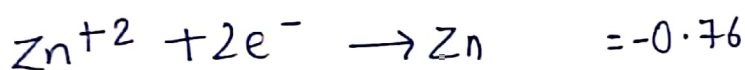


0.76 volt



0.80 volt

Sol



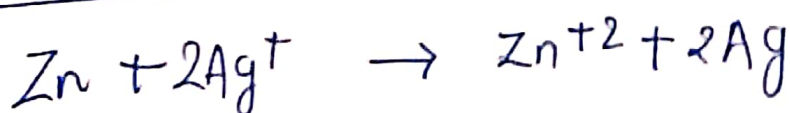
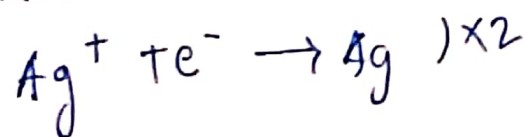
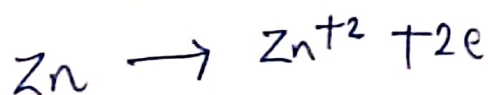
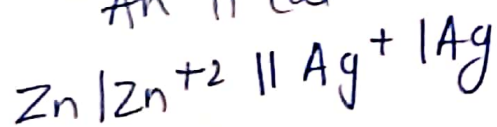
$$E^{\circ} = E_{\text{cat}} - E_{\text{an}}$$

$$E^{\circ} = +ve$$

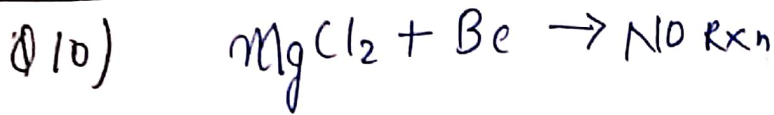
$$E^{\circ} = 0.80 - (-0.76)$$

$$E^{\circ} = 1.56 \text{ volt}$$

An || cat

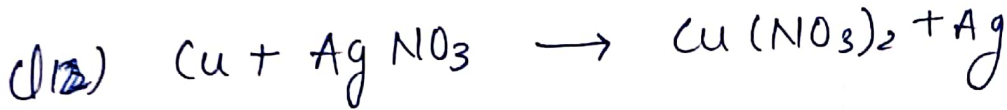


10	6	25
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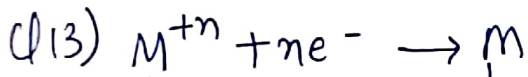
Q11) 0.00 volt

(Ans)



Cu^+ $\rightarrow$ colourless

$\text{Cu}^{+2} \rightarrow$ colour Full
(Blue)



red

↓ red potential

$$m^{+2}/M \rightarrow \text{Reduction}$$

Value $\rightarrow$ Max = -ve

$$M_1 = -0.30$$

$$M_2 = -3.05$$

$$M_3 = -1.66 \text{ V}$$

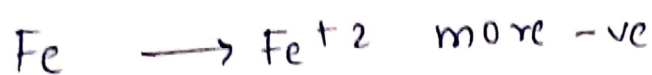
$$m_2 > m_3 > m_1$$

Q14) A

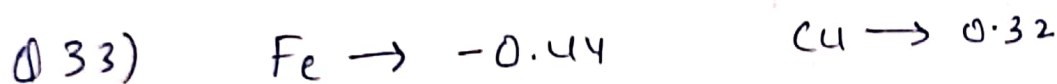
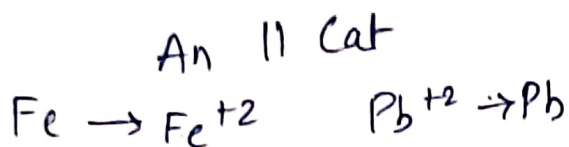
Q(5) Hg

Q 19) $Ca > K > Li$

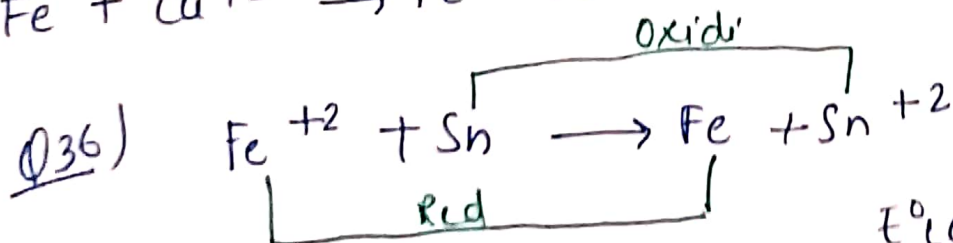
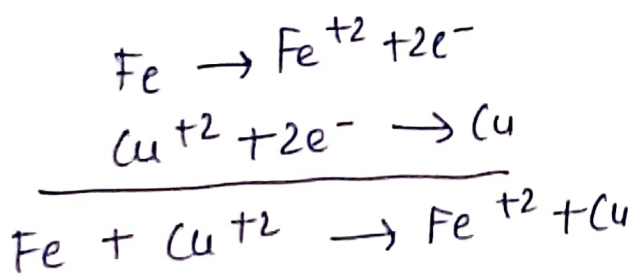
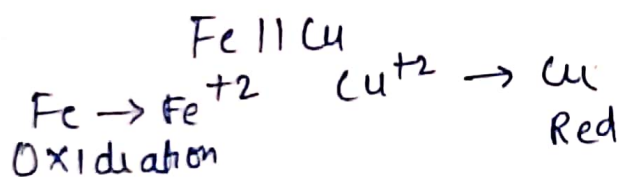
Q20) Opt (1)



$$E^{\circ}_{cell} = (\text{less -ve})_{cat} - (\text{-more -ve})_{An}$$



$$E^{\circ}_{cell} = +ve = E_{cat} - E_{an} = 0.32 - (-0.44)$$

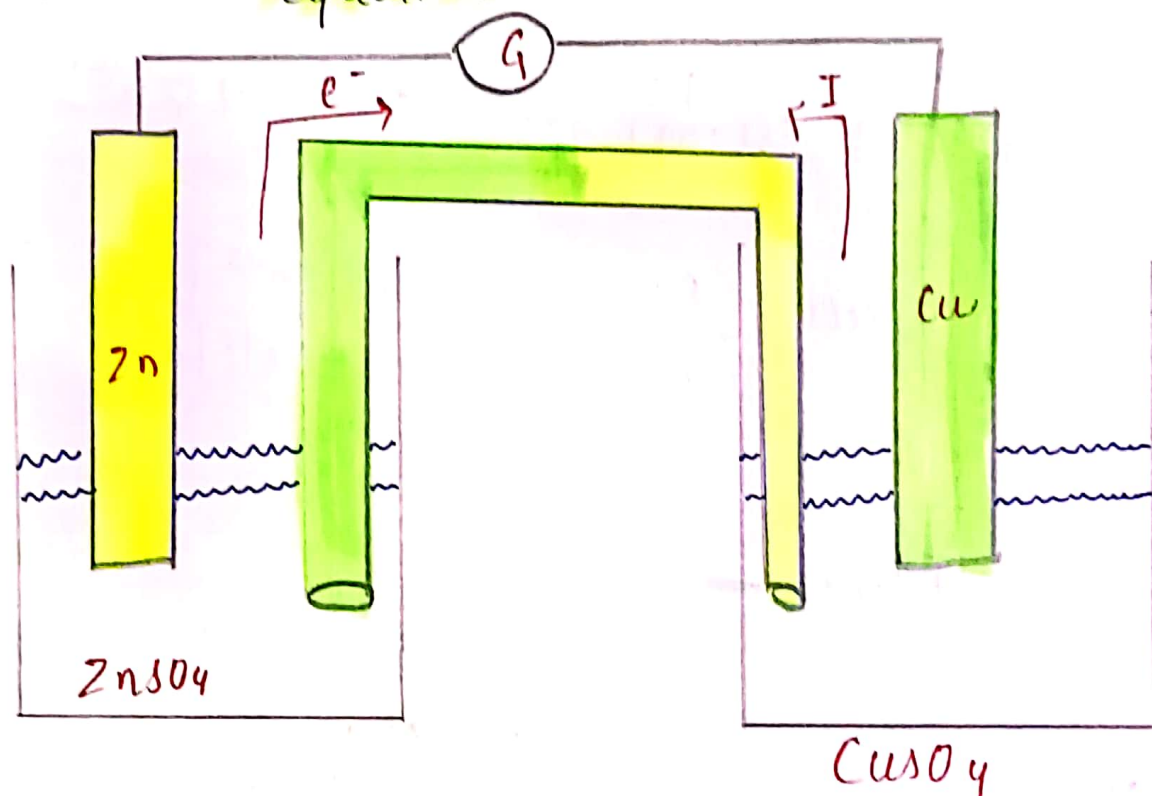


$$\begin{aligned} E^{\circ}_{cell} &= E^{\circ}_{Red} - E^{\circ}_{ox} \\ E^{\circ}_{cell} &= -0.40 + 0.14 \\ E^{\circ}_{cell} &= -0.30 \text{ volt} \end{aligned}$$

Q37) $E^{\circ}_{cell} = +ve$
 $= 0.34 + 2.37$
 $= 2.71 \text{ volt}$

* NERST Equation

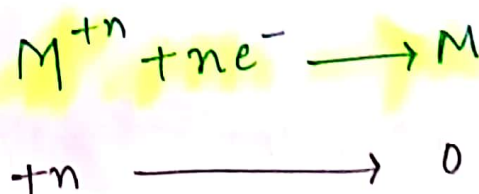
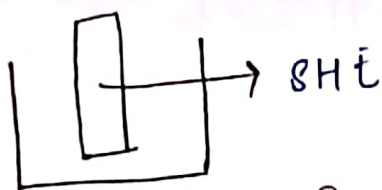
Danrell cell



"What is NERST Equation"

"Relationship between E_{cell} & concentration"

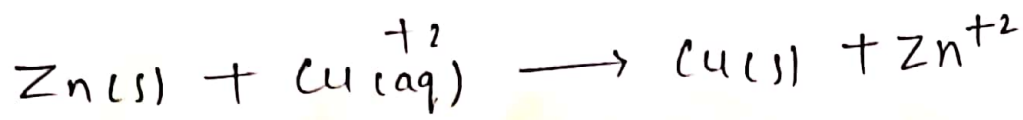
E° - Standard



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

"NERST EQUATION FOR HALF CELL"

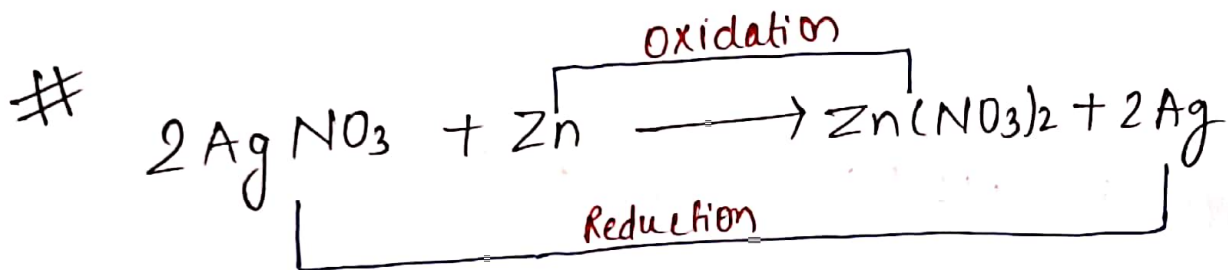
FOR DANIEL CELL



$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{RT}{nf} \ln \frac{[\text{Cu}][\text{Zn}^{+2}]}{[\text{Zn}][\text{Cu}^{+2}]}$$

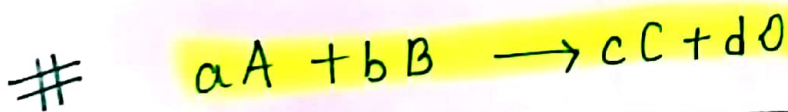
Metal/Solid
↓
unity

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} = \frac{RT}{nf} \ln \frac{[\text{Zn}^{+2}]}{[\text{Cu}^{+2}]}$$

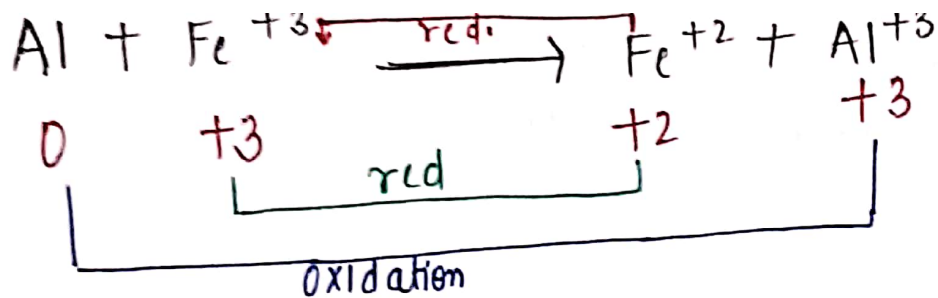


$$E_{\text{cell}} = (E^{\circ}_{\text{Ag}^+/\text{Ag}} - E^{\circ}_{\text{Zn}^{+2}/\text{Zn}}) - \frac{RT}{nf} \ln \frac{[\text{Zn(NO}_3)_2]}{[\text{AgNO}_3]^2}$$

$$E_{\text{cell}} = E^{\circ}_{\text{Ag}^+/\text{Ag}} - E^{\circ}_{\text{Zn}^{+2}/\text{Zn}} - \frac{0.0591}{n} \log_{10} \frac{[\text{Zn(NO}_3)_2]}{[\text{AgNO}_3]^2}$$



$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0591}{n} \log_{10} \frac{[\text{C}]^c [\text{D}]^d}{[\text{A}]^a [\text{B}]^b}$$



STEP - I

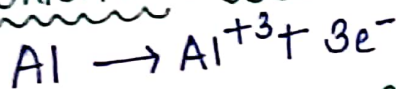
① Calculation of E°_{cell}

$$E^\circ_{\text{cell}} = E^\circ_{\text{red}} - E^\circ_{\text{ox}}$$

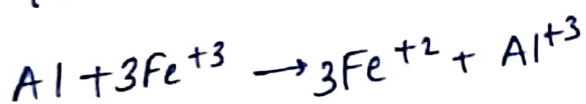
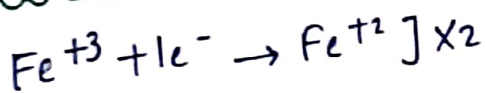
$$E^\circ_{\text{cell}} = E^\circ_{\text{Fe}^{+3}/\text{Fe}^{+2}} - E^\circ_{\text{Al}^{+3}/\text{Al}}$$

STEP - II

OXIDATION HALF REACTION



REDUCTION HALF REACTION



STEP - IV

calculation of log

$$\log \begin{cases} \rightarrow \ln = \log_e \\ \rightarrow \log_{10} \end{cases}$$

$$2.303 \log_{10} x = \ln x$$

STEP - II

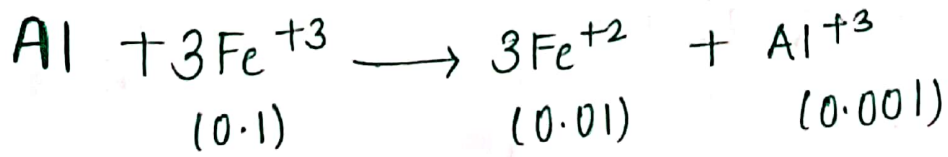
Calculation of n

$$n = 3$$

STEP - III

$\frac{\text{Product}}{\text{Reactant}}$

$$\frac{[\text{Al}^{+3}] [\text{Fe}^{+2}]^3}{[\text{Fe}^{+3}]^3}$$



$$E_{\text{cell}} = E^{\circ}_{\text{Fe}^{+3}/\text{Fe}^{+2}} - E^{\circ}_{\text{Al}^{+3}/\text{Al}} - \frac{0.0591}{3} \log_{10} \frac{[\text{Al}^{+3}][\text{Fe}^{+2}]^3}{[\text{Fe}^{+3}]^3}$$

* Calculation of $\frac{RT}{nF} \ln$ *

0.0821

$$\frac{RT}{nF} \ln$$

$\xrightarrow{298\text{ K}}$
 $\xrightarrow{2.303 \log_{10}}$
 $\hookrightarrow 96500$

1 mole e^- flow

$$1 \times 6.02 \times 10^{23} \times 1.6 \times 10^{-19}$$

$$= 96480 \text{ C}$$

$$= \boxed{96500 \text{ C} = 1 \text{ F}}$$

Example For Daniel cell

$$E^\circ_{\text{Zn}^{+2}/\text{Zn}} = -0.763 \text{ volt}$$

$$E^\circ_{\text{Cu}^{+2}/\text{Cu}} = 0.337 \text{ volt}$$

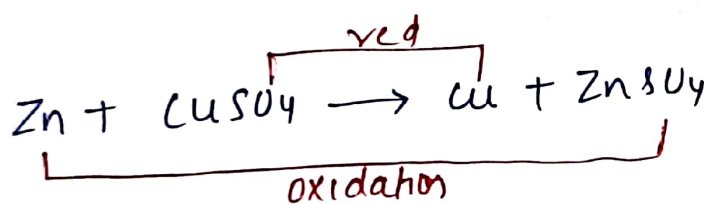
$$\text{Zn}^{+2} = 0.01 \quad \text{Cu}^{+2} = 0.001$$

Step-I E°_{cell} calculation

$$E^\circ_{\text{cell}} = E^\circ_{\text{red}} - E^\circ_{\text{ox}}$$

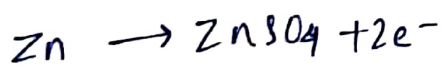
$$E^\circ_{\text{cell}} = 0.337 + 0.763$$

$$E^\circ_{\text{cell}} = 1.10 \text{ volt}$$

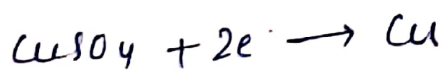


Step-II calculation of n

Oxidation half Reaction



Reduction half Reaction



$$n = 2$$

STEP-III concentration find

$$\frac{\text{Product}}{\text{Reactant}} = \frac{[\text{ZnSO}_4]}{[\text{CuSO}_4]}$$

STEP-IV log calculation

$$E^\circ_{\text{cell}} = E_{\text{cell}} - \frac{0.0591}{n} \log \frac{P}{R}$$

$$E_{\text{cell}} = 1.10 - \frac{0.0591}{2} \log_{10} \frac{0.01}{0.001}$$

$$E_{\text{cell}} = 1.1 + \frac{0.06}{2} \log_{10} 10$$

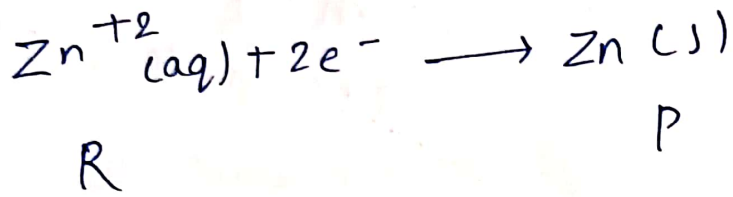
$$E_{\text{cell}} = 1.1 + 0.03$$

$$\boxed{E_{\text{cell}} = 1.13 \text{ volt}}$$

Pg = 34

DPP-07

Ques. 1 Zn^{+2} / Zn



$$n = 2$$

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0591}{2} \log_{10} \frac{1}{Zn^{+2}}$$

$$\underset{P}{1} \rightarrow 0.01$$
$$\underset{P}{R} = 10^{-2}$$

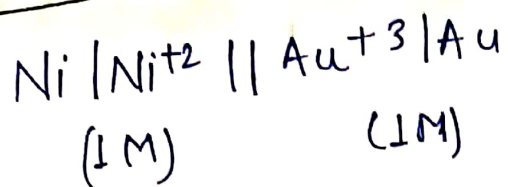
$$E_{cell} = E^{\circ}_{cell} - \frac{0.0591}{2} \log \frac{1}{10^{-2}}$$

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0591}{2} \times 2 \log 10$$

$$[E_{cell} = \cancel{no.} E^{\circ}_{cell} - 0.0591 \text{ volt}]$$

⊙

Q.2)



$$\text{Ni}^{+2} | \text{Ni} \rightarrow -0.25 \text{ volt}$$

$$\text{Au}^{+3} | \text{Au} \rightarrow 1.50 \text{ volt}$$

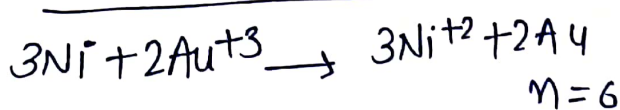
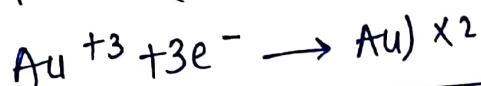
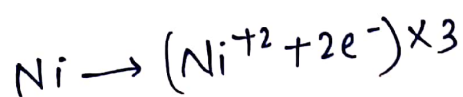
Step-I $E^{\circ}_{\text{cell}} = E_{\text{Right}} - E_{\text{Left}}$

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{Au}^{+3} | \text{Au}} - E^{\circ}_{\text{Ni}^{+2} | \text{Ni}}$$

$$E^{\circ}_{\text{cell}} = 1.50 + 0.25$$

$$E^{\circ}_{\text{cell}} = 1.75 \text{ Volt}$$

Step-II



Step-III

$$\frac{P}{R} = \frac{[\text{Ni}^{+2}]^3}{[\text{Au}^{+3}]^2}$$

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0591}{6} \log_{10} \frac{[\text{Ni}^{+2}]^3}{[\text{Au}^{+3}]^2}$$

Step-IV

$$E_{\text{cell}} = E^{\circ}_{\text{cell}}$$

$$E_{\text{cell}} = 1.75 \text{ Volt}$$

17/06/2025

Relationship Between ΔG , ΔG° , E°_{cell} , E_{cell} , K_c

K_c

ΔG :- Gibbs Free Energy

+ve
Non Spontaneous

-ve
Spontaneous
अपेक्षा

Zero
Equilibrium

$$\Delta G = -nFE_{\text{cell}}$$

$$\Delta G = -nF \left(E^\circ_{\text{cell}} - \frac{RT}{nF} \ln \frac{P}{R} \right)$$

If we are in equilibrium

$$\Delta G = 0 \quad \frac{P}{R} = K_c$$

$$0 = -nF \left(E^\circ_{\text{cell}} - \frac{RT}{nF} \ln K_c \right)$$

$$= -nFE^\circ_{\text{cell}} + \frac{nFRT}{nF} \ln K_c$$

$$nF E^\circ_{\text{cell}} = RT \ln K_c$$

$$E^\circ_{\text{cell}} = \frac{RT}{nF} \ln K_c$$

$$E^{\circ}_{\text{cell}} = \frac{2.303 RT}{n F} \log_{10} K_c$$

$$E^{\circ}_{\text{cell}} = \frac{0.0591}{n} \log_{10} K_c$$

$$\Delta G^{\circ} = -n F E^{\circ}_{\text{cell}}$$

$$\Delta G^{\circ} = -RT \ln K_c$$

$$\Delta G^{\circ} = -2.303 RT \log_{10} K_c$$

Total Relationships

$$\rightarrow \Delta G = -n F E_{\text{cell}}$$

$$\rightarrow \Delta G^{\circ} = -n F E^{\circ}_{\text{cell}}$$

$$\rightarrow \Delta G^{\circ} = -2.303 RT \log_{10} K_c$$

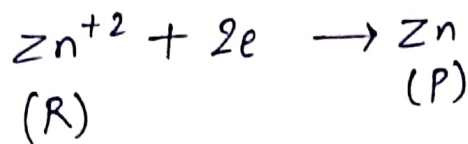
$$\rightarrow E^{\circ}_{\text{cell}} = \frac{2.303 RT}{n F} \log_{10} K_c$$

$$\rightarrow E^{\circ}_{\text{cell}} = \frac{0.0591}{n} \log_{10} K_c$$

21/06/2025

DPP-07

Ques. 1

 $\hookrightarrow 0.01 \text{ M}$

$$E_{\text{Zn}^{+2}/\text{Zn}} = E^{\circ}_{\text{Zn}^{+2}/\text{Zn}} - \frac{0.0591}{n} \log_{10} \frac{P}{R}$$

$$m = \frac{\text{mol}}{\text{Volume}} = \frac{\text{mol}}{100} = 0.01 \text{ M}$$

$$E_{\text{Zn}^{+2}/\text{Zn}} = E^{\circ}_{\text{Zn}^{+2}/\text{Zn}} - \frac{0.0591}{2} \log_{10} \frac{1}{10^{-2}}$$

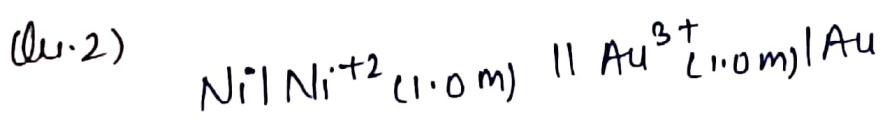
$$E_{\text{Zn}^{+2}/\text{Zn}} = E^{\circ}_{\text{Zn}^{+2}/\text{Zn}} - \frac{0.0591 \times 2}{2} \log_{10} 10$$

$$E_{\text{Zn}^{+2}/\text{Zn}} = E^{\circ}_{\text{Zn}^{+2}/\text{Zn}} - 0.0591$$

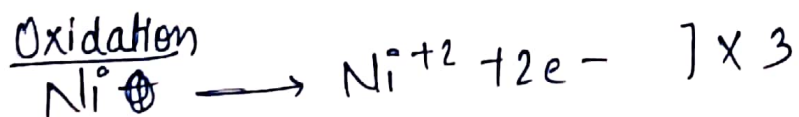
$$E_{\text{Zn}^{+2}/\text{Zn}} = E^{\circ}_{\text{Zn}^{+2}/\text{Zn}} - 0.0591 \times 10^3$$

So, opt (2) Decreases by 59 mv.

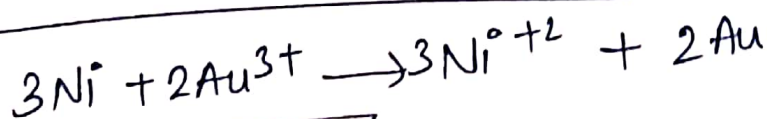
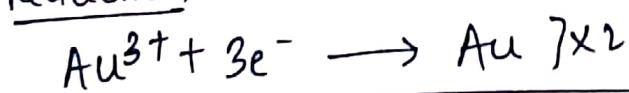
Ques. 2)



Oxidation



Reduction



$$\boxed{n=6}$$

$$E^{\circ}_{\text{cell}} = E_{\text{Right}} - E_{\text{Left}}$$

$$E^{\circ}_{\text{cell}} = 1.50 + 0.25 = 1.75 \text{ Volt}$$

Step - III

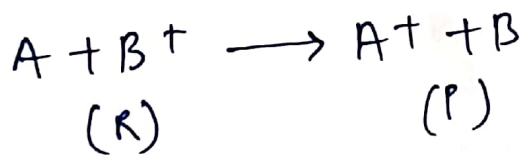
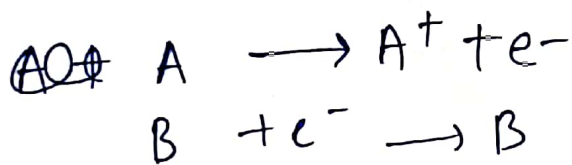
Concentration

$$\frac{P}{R} = \frac{[Ni^{+2}]^3 [Au]^2}{[Au^{+3}]^2}$$

$$E_{cell} = E^{\circ}_{cell} = \frac{0.0591}{n} \log_{10} \frac{P}{R}$$

$$E_{cell} = 1.75 - \frac{0.0591}{6} \log_{10} 1$$

$$E_{cell} = 1.75 \text{ Volt}$$



$$E^{\circ}_{cell} = E^{\circ}_B - E^{\circ}_A$$

$$E^{\circ}_{cell} = 0.75 - 0.5$$

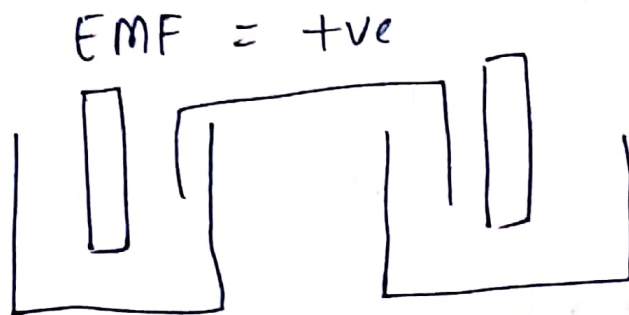
$$E^{\circ}_{cell} = 0.25$$

$$n = 1$$

$$\frac{P}{R} = 1$$

$$E_{cell} = 0.25 \text{ Volt}$$

24

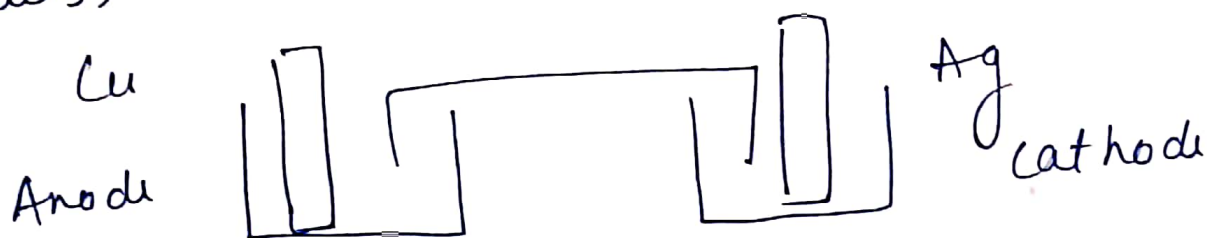


$$E^{\circ}_{\text{cell}} = E_R - E_L$$

$$= 0.33 + 2.37$$

$$E^{\circ}_{\text{cell}} = 2.70 \text{ volt}$$

Ques 5)



$$E^{\circ}_{\text{cell}} = +ve$$

$$E^{\circ}_{\text{cell}} = 0.46 \text{ volt}$$

Ques 10

$$E^{\circ}_{\text{cell}} = ?$$

$$E_{\text{cell}} = 0.5105 \text{ Volt}$$

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0591}{n} \log_{10} \frac{P}{R}$$

$$n = 2$$



(R)

(P)

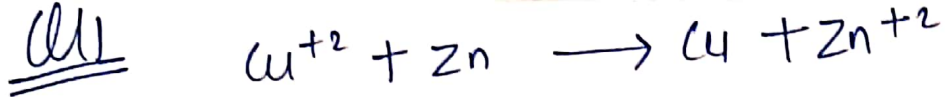
(1)

(10)

$$0.5105 = E^{\circ}_{\text{cell}} - \frac{0.0591}{2} \log_{10} \frac{10}{1}$$

$$E^{\circ}_{\text{cell}} = 0.5105 + 0.03$$

$$E^{\circ}_{\text{cell}} = 0.5405 \text{ Volt}$$



case I

	(R)	(P)
	1	0.01

$$E_{\text{cell}_1} = E^{\circ}_{\text{cell}} - \frac{0.0591}{2} \log 10^{-2}$$

$$E_{\text{cell}_1} = E^{\circ}_{\text{cell}} = +0.591$$

$$E_{\text{cell}_1} = 1.10 + 0.0591$$

$$E_{\text{cell}_1} =$$

case - II) P = 1 R = 0.01

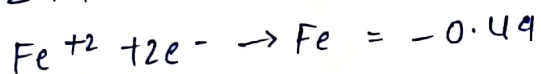
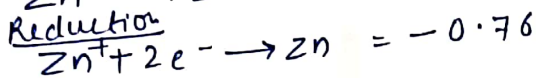
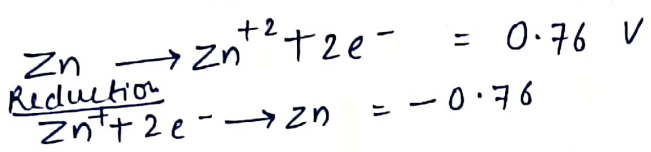
$$E_2 (E_{\text{cell}})_2 = E^{\circ}_{\text{cell}} - \frac{0.0591}{2} \log_{10} \frac{1}{10^{-2}}$$

$$(E_{\text{cell}})_2 = E^{\circ}_{\text{cell}} - 0.0591$$

$$(E_{\text{cell}})_2 = 1.10 - 0.0591$$

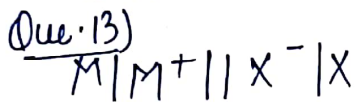
$$\boxed{E_1 > E_2}$$

Qw.12 $E \rightarrow$ oxidation



$$E^\circ_{\text{cell}} = -0.44 + 0.76 = 0.32 \text{ volt}$$

Que.13)

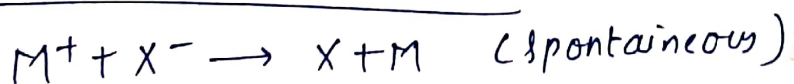
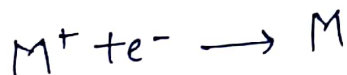
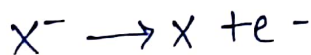
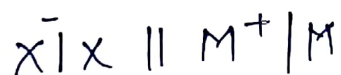


$$E^\circ_{\text{cell}} = +ve$$

$$E^\circ_{\text{cell}} = 0.33 - 0.44$$

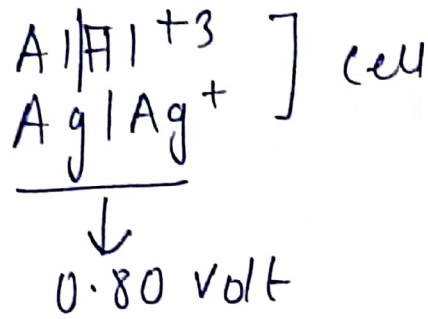
$$E^\circ_{\text{cell}} = -0.11$$

Due to $E^\circ_{\text{cell}} = -ve$ we change electrode



Qw.14 (1) Yes

Qw.15)

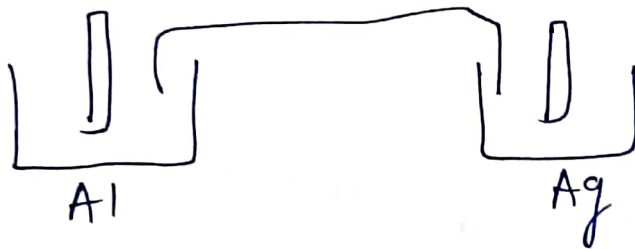


$$E_{\text{cell}} = 2.46 \text{ Volt}$$

$$2.46 = 0.80 - E^\circ \text{Al}^{3+}/\text{Al}$$

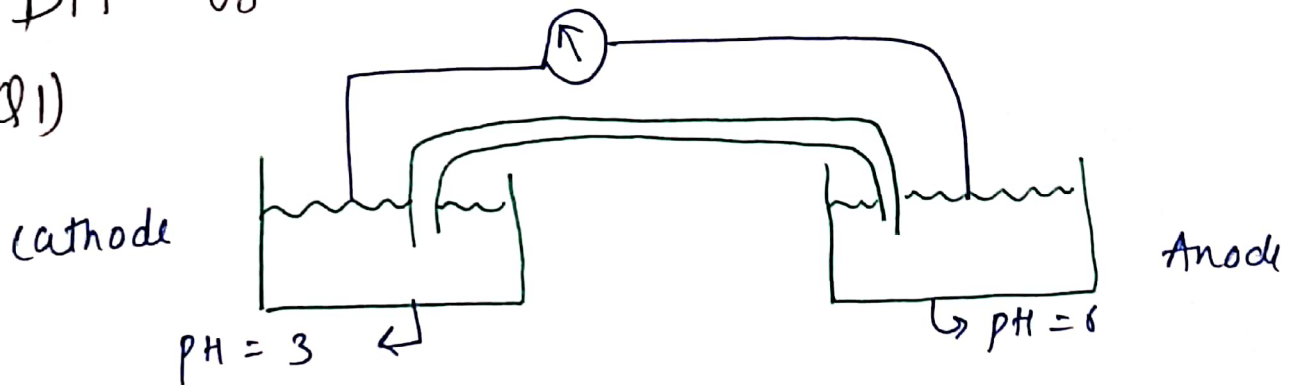
$$E^\circ \text{Al}^{3+}/\text{Al} = 0.80 - 2.46$$

$$E^\circ \text{Al}^{3+}/\text{Al} = -1.66 \text{ Volt}$$



DPP - 08

Q1)



$$E_{\text{H}^+/\text{H}_2} = -0.0591 \times \text{pH}$$

$$E_{\text{H}^+/\text{H}_2} = -0.06 \times 3$$

$$E_{\text{H}^+/\text{H}_2} = -0.18$$

$$E_{\text{H}^+/\text{H}_2} = -0.0591 \times 6$$

$$E_{\text{H}^+/\text{H}_2} = -0.36$$

$$E^\circ_{\text{cell}} = +ve$$

$$E^\circ_{\text{cell}} = -0.18 + 0.36$$

$$E^\circ_{\text{cell}} = +0.18 \text{ Volt}$$

* Concept of ΔG

* Couple Reaction

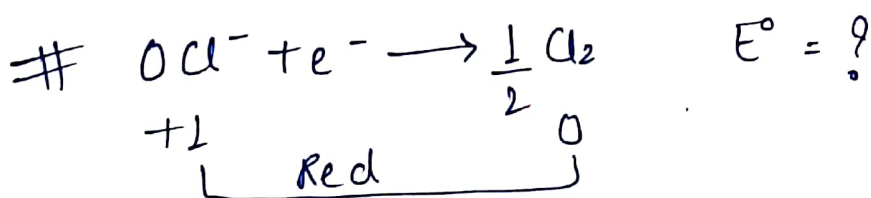
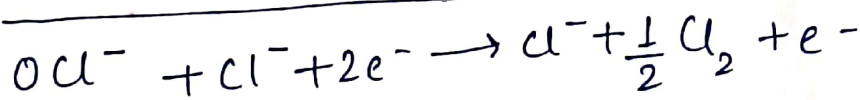
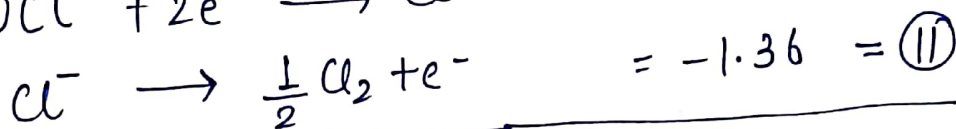
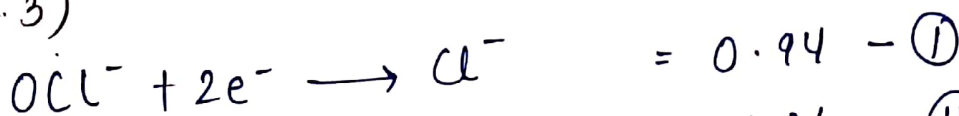
* Disproportionation Reaction

ΔG is Extensive property i.e. it depends upon number of moles and E°_{cell} is intensive property that means it doesn't depend upon concentration or mole.

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

DPP-08

Q. 3)



Eq (I) + Eq (II)

$$\Delta G_1^\circ = -2 \times F \times 0.94$$

$$\Delta G_2^\circ = +1 \times F \times 1.36$$

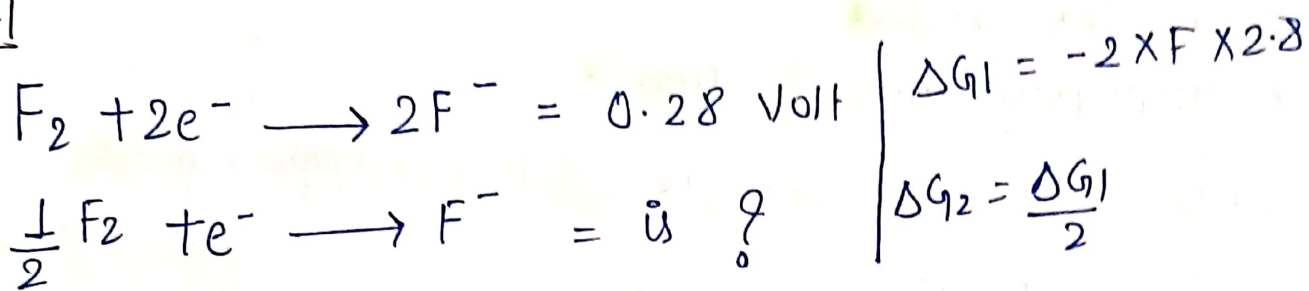
$$\Delta G_3 = -2 \times 0.94 \times F + 1 \times F \times 1.36$$

$$\Delta G_3 = -0.52 F$$

$$-1 \times F \times E^\circ_{\text{cell}} = -0.52 F$$

$E^\circ_{\text{cell}} = 0.52 \text{ Volt}$

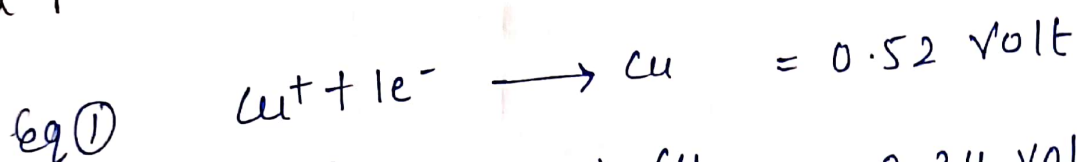
Ques. 1



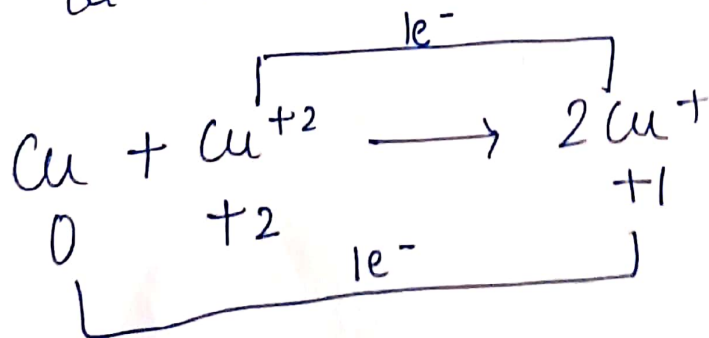
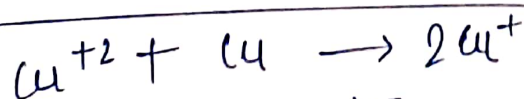
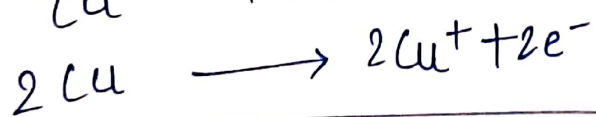
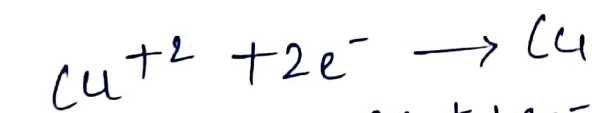
$$= -1 \times F \times E^\circ_{\frac{1}{2} F_2/F} = -F \times 2.8$$

$$E^\circ_{\frac{1}{2} F_2/F} = 2.8 \text{ Volt}$$

Ques. 7



$$\text{Eq (1)} - (\text{Eq (2)}) \times 2$$



~~total~~ $\Delta G_2 - 2 \times \Delta G_1 = \Delta G_3$

$$-2 \times F \times 0.34 + 2 \times 0.52 \times F = 0.43$$

$$\Delta G_3 = -nFE^\circ_{Rxn} = 2F(0.52 - 0.34)$$

$$-1 \times F \times E^{\circ}_{\text{rxn}} = F \times 0.36$$

DPP - 08

$$\text{Cu} + \underset{\text{(R)}}{2\text{Ag}^+} \longrightarrow \underset{\text{(P)}}{\text{Cu}^{+2}} + 2\text{Ag}$$

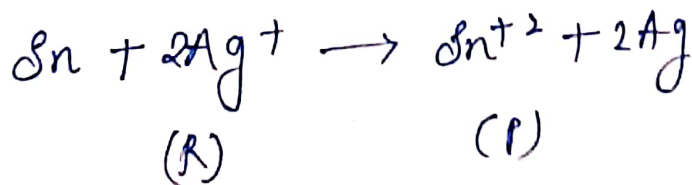
$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0591}{2} \log_{10} \frac{P}{R} = \frac{Cu^{+2}}{[Ag^{+}]^2}$$

$p = Cu^+$ (↑) so, $E_{cell} = \text{Decrease}$

$$R = Ag^+(\uparrow) \quad = \quad E_{cell} = \text{Increases}$$

Opt (J)

Q5)

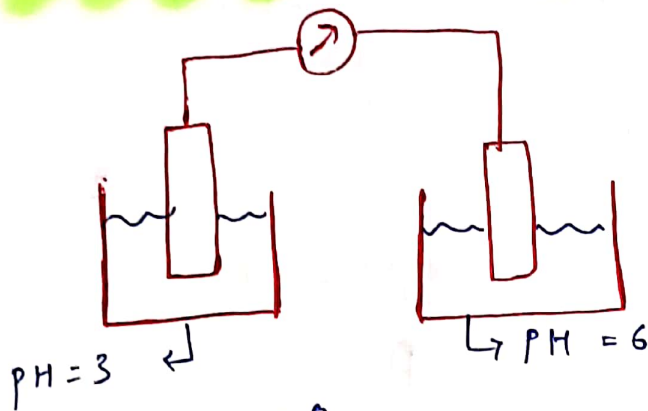

$$\left. \begin{array}{l} \text{Sn}^{+2} = P \downarrow \\ \text{Ag}^+ = R \uparrow \end{array} \right\} \text{For } \varepsilon_{\max} \text{ opt (L)}$$

SO Starts
Special Case

Concentration of cell

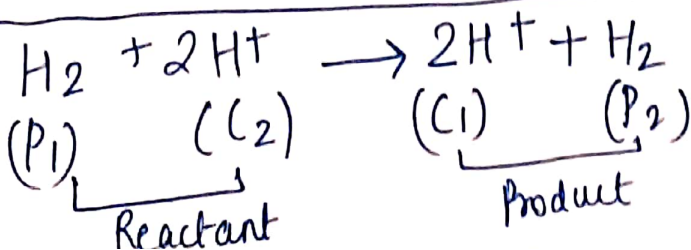
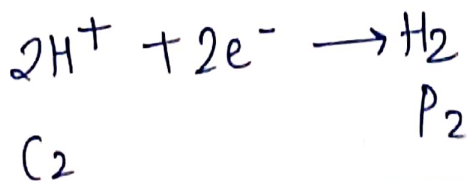
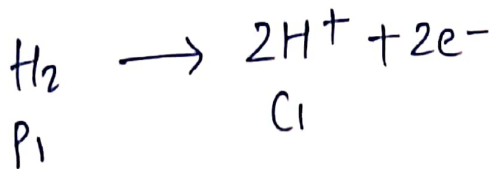
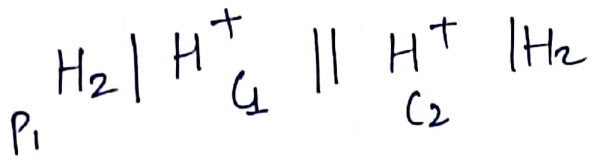
A concentration cell is a form of galvanic cell that has 2 equivalent half cell of the same material having difference only in the concentration such cell. Because of cathode, anode same so E°_{cell} is equal to zero.

Ex -



$$E_{\text{cell}} = \cancel{E^\circ_{\text{cell}}} = 0 - \frac{RT}{nF} \ln \frac{P}{R}$$

Ques. 6



$$E_{\text{cell}} = E_{\text{cell}}^{\circ} = 0$$

$$E_{\text{cell}} = -\frac{RT}{2F} \ln \left(\frac{(C_1)^2}{(C_2)^2} \cdot \frac{P_2}{P_1} \right)$$

$$E_{\text{cell}} = -\frac{RT}{2F} \ln \left(\left(\frac{C_1}{C_2} \right)^2 \left(\frac{P_2}{P_1} \right) \right)$$

When pressure are same

$$E_{\text{cell}} = -\frac{RT}{2F} \ln \left(\frac{C_1}{C_2} \right)^2$$

$$E_{\text{cell}} = -\frac{RT}{F} \ln \left(\frac{C_1}{C_2} \right)^{2 \times \frac{1}{2}}$$

$$E_{\text{cell}} = -\frac{RT}{F} \ln \frac{C_1}{C_2}$$

When concentration are same

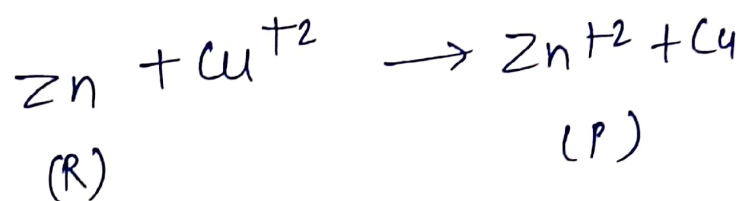
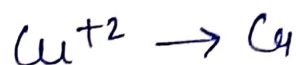
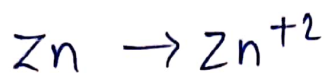
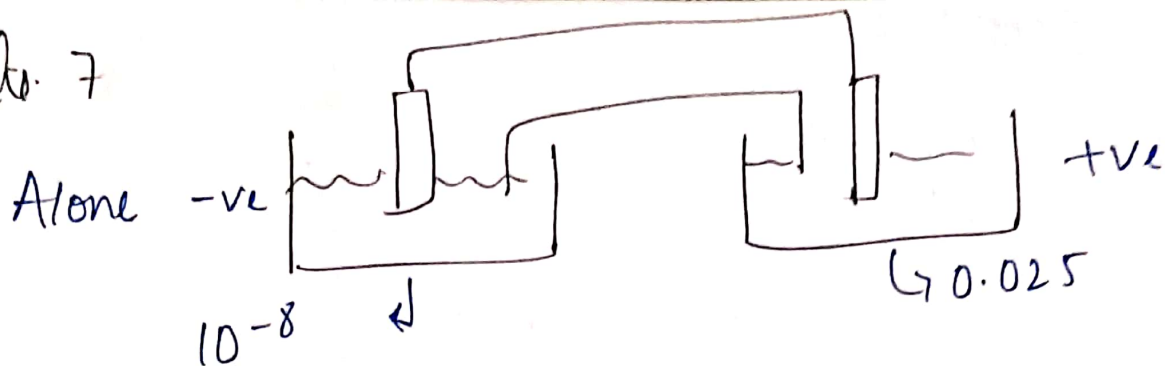
$$E_{\text{cell}} = -\frac{RT}{2F} \ln \left(\frac{C_1}{C_2} \right)^2 \frac{P_2}{P_1}$$

$$E_{\text{cell}} = -\frac{RT}{2F} \ln \frac{P_2}{P_1}$$

$$E_{\text{cell}} = \frac{RT}{2F} \ln \left(\frac{P_2}{P_1} \right)^{-1}$$

$$E_{\text{cell}} = \frac{RT}{2F} \ln \left(\frac{P_1}{P_2} \right)$$

Q. 7



$$\frac{P}{R} = \frac{10^{-8}}{0.025} =$$

$$E_{\text{cell}} = \frac{-0.0591}{1} \log_{10} \frac{1000 \times 10^{-8}}{25}$$

$$E_{\text{cell}} = -0.0591 \log 2^2 + \log 10^{-7}$$

$$E_{\text{cell}} = -0.0591 (0.60 - 7)$$

$$E_{\text{cell}} = +0.06 \times 6.4$$

$$[E_{\text{cell}} = 0.384 \text{ volt}]$$

Qw. 8)

(0.1) HCl = Acid $\rightarrow$ Strong Acid

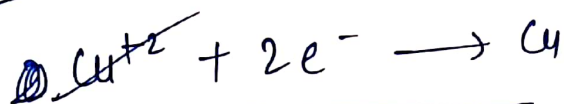
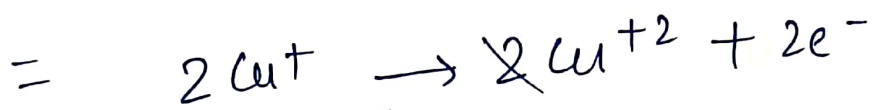
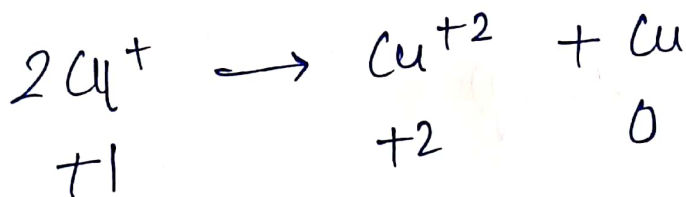
(0.1) CH₃COOH = Acid $\rightarrow$ Weak Acid

$$pH_{HCl} \neq pH_{CH_3COOH}$$

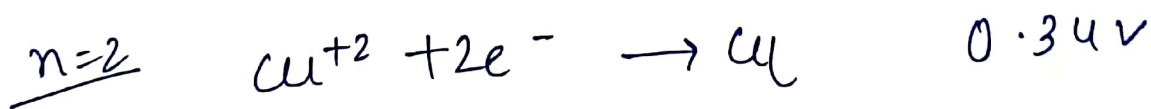
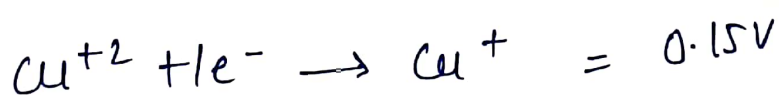
opt ① $E_{cell} \neq 0$

Qw. 9

n=2



n=1



$$- 2 \times \Delta G_1 + \Delta G_2 = \Delta G_3$$

$$- 2 \times 1 \times F \times 0.15 + (-2 \times F \times 0.34) = -2 \times F \times E$$

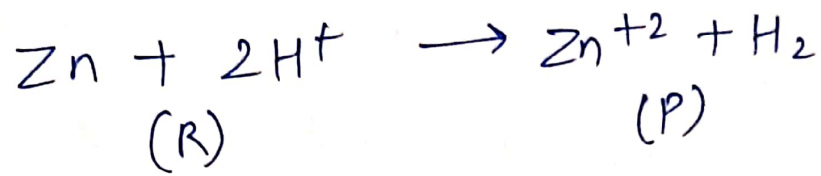
$$+ 0.15 - 0.34 = -E$$

$$- 0.19 = -E$$

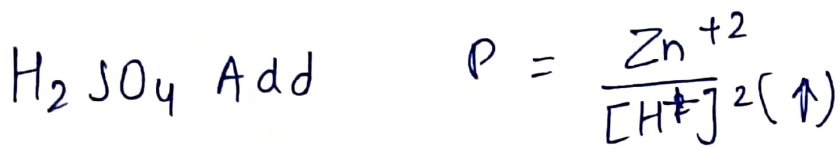
opt (4) .

$$E_{cell}^{\circ} = 0.19 \text{ Volt}$$

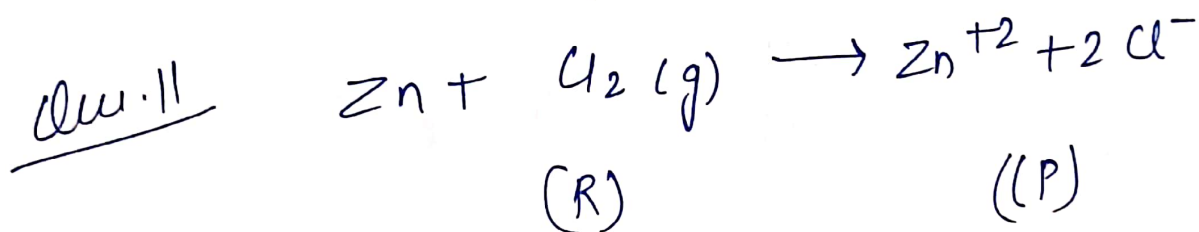
Ques 10



R(↑) → Right forward



Ecell (↑) Opt(1)



$$\frac{P}{R} = \frac{\downarrow [\text{Zn}^{+2}] [\text{Cl}^-]^2 \downarrow}{P\text{Cl}_2(1)}$$

Opt (2) [Zn²⁺] should be decreased.

24/6/25

ELECTROLYTIC CONDUCTION (25 Stars)

Electrolytic Conduction

→ Ohm's law

Voltage is directly proportional to the current.

$$V \propto R$$

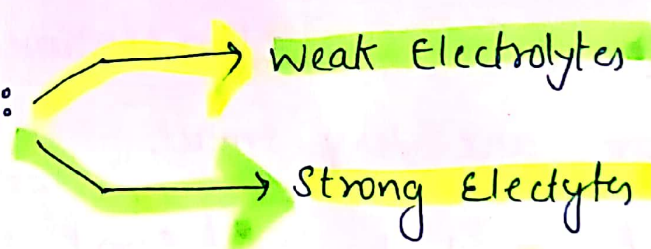
$$V = IR$$

↳ R is called Resistance.

* Its unit : Ω (ohm)

In the Electrochemistry we study only about flow of ions which is only possible when we add electrolytes in water.

Electrolytes are of 2 types :



- Weak Electrolytes
- Strong Electrolytes

Weak Electrolytes are substances which do not dissociate 100%. They likely dissolve 5-10%. [Sugar, Alcohol] etc.

Strong Electrolytes are completely dissociated in the presence of water.

Ex - NaCl, NaBr, CaCl₂,

→ Conductance

It is the property of conductor which favours flow of current.

It is the reciprocal of Resistance.

It's unit are ohm^{-1} or siemens. $[\Omega^{-1}]$ / $[S]$ or $[\text{V}] \text{mho}^{-1}$.

→ Specific Resistance / Resistivity

The Resistance of a conductor is directly proportional to an length and inversely proportional to Cross Section Area.

It's unit are $[\Omega \text{cm}]$ ohm centimeter, Because length of conductor are very small.

→ Specific Conductance / Conductivity

It is defined as the reciprocal of specific Resistance.

It's unit's are $[\text{ohm}^{-1} \text{cm}^{-1}]$ $[\Omega^{-1} \text{cm}^{-1}]$.

Specific Conductance equal to conductance $\times$ cell constant

Cell constant is denoted by $[G^*]$. which is equal to $\frac{l}{A}$. and its unit are $[\text{cm}^{-1}]$

$$\left[K = \frac{1}{R} \cdot \frac{l}{A} = G \times G^* \right]$$

→ Molar conductivity

It is defined as the conductance of all the ions produced by one ^{gram eq.} ~~mole~~ _{mole} of electrolytes present in the given volume of solution.

It is denoted by Λ_m and μ .

It's unit are $[\text{ohm}^{-1} \text{cm}^2 \text{mol}^{-1}]$.

~~Equivalent~~

→ Equivalent conductivity

It is defined as the conductance of all the ions are produced by one gram equivalent

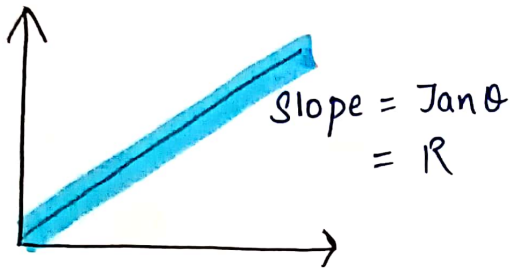
It's unit is $[\text{ohm}^{-1} \text{cm}^2 \text{g eq.}^{-1}]$.

It is denoted by Λ_{eq} .

* Ohm's law

$$V \propto R$$

$$V = IR$$



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

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

$$\rho = \frac{\Omega \text{ m}^2}{\text{m}}$$

$$\rho = \Omega \text{ m}$$

$$\rho = \Omega \text{ cm}$$

R = Resistance
Unit = Ω , ohm

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

ρ = Resistivity

l = length of electrode

A = Area of cross section
of electrode = πr^2

$$\frac{1}{R} = \text{Conductance } [G]$$

$$\frac{1}{R} = G = \frac{\Omega^{-1}}{\text{or Siemens}}$$

$$\frac{1}{\rho} = \text{Kappa } [K]$$

$$K = \Omega^{-1} \text{ cm}^{-1} [\Omega^{-1} \text{ cm}^{-1}]$$

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

$$K = G \times G^*$$

$$R = \frac{1}{K} \times \frac{l}{A}$$

$$\frac{l}{A} = \frac{\text{cm}}{\text{cm}^2} = \text{cm}^{-1}$$

$$G^* = \text{cm}^{-1}$$

$$K = \frac{1}{R} \times \frac{l}{A}$$

$$K = G \times \left(\frac{l}{A} \right) \rightarrow \text{cell constant}$$

~~Symbol~~

Ω , R = Resistance

$\Omega \text{ cm}$, ρ = Resistivity, Specific resistance

Ω^{-1} , G = conductance

$\Omega^{-1}, \text{cm}^{-1}$, K = conductivity, Specific conductance

$\Omega^{-1} \text{ mol}^{-1} \text{ cm}^2$, Λ_m = molar conductivity

$\Omega^{-1} \text{ eq}^{-1} \text{ cm}^2$, Λ_{eq} = Equivalent conductivity

cm^{-1} , G^* = cell constant

$$R = \frac{\rho l}{A} \left[\begin{array}{l} \xrightarrow{l \text{ cm}} \\ \xrightarrow{A \text{ cm}^2} \end{array} \right] \text{unit}$$

$R = \rho$ = Specific Resistance

[K] conductivity

Molar conductivity
[Λ_m]

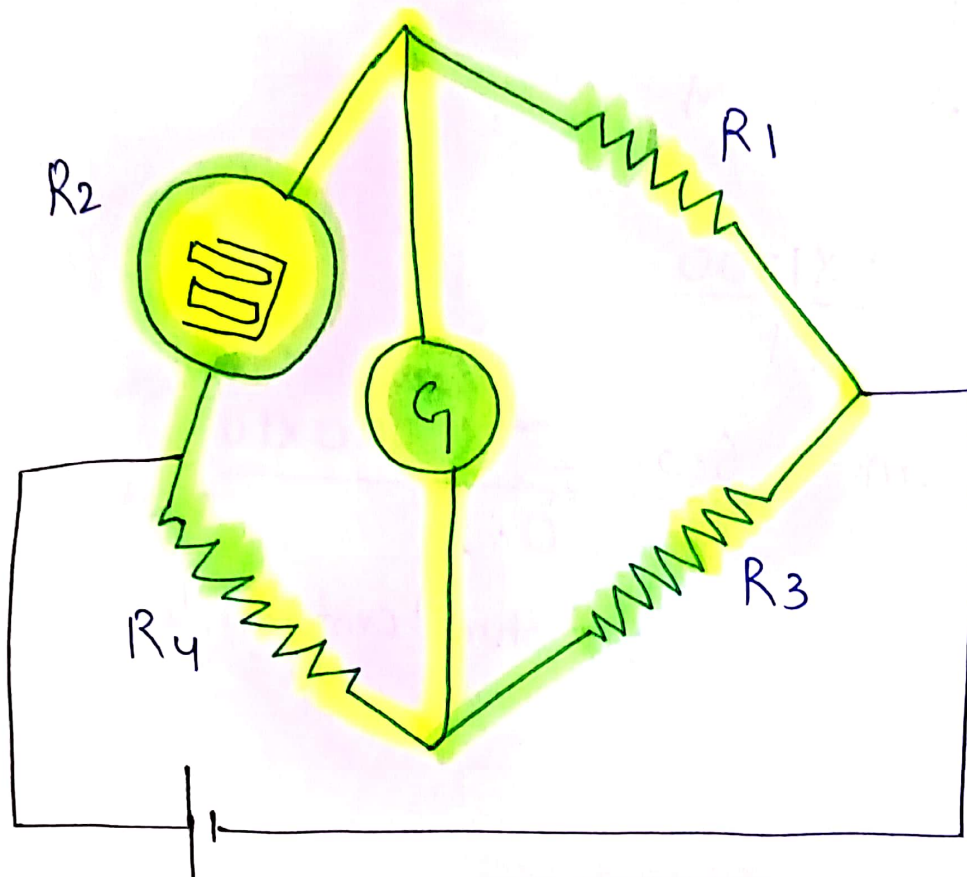
Equivalent
conductivity
[Λ_{eq}]

Λ_m = molar conductivity

$$\Lambda_m = \frac{K \times 1000}{M} = \frac{\text{ohm}^{-1} \text{cm}^{-1}}{\frac{\text{mol}}{\text{cm}^3}} = \text{ohm}^{-1} \text{mol}^{-1} \text{cm}^2$$

$$\Lambda_{eq} = \frac{K \times 1000}{N} = \frac{\text{ohm}^{-1} \text{cm}^{-1}}{\frac{\text{geq}}{\text{cm}^3}} = \text{ohm}^{-1} \text{geq}^{-1} \text{cm}^2$$

WHEATSTONE BRIDGE



$$\frac{R_2}{R_4} = \frac{R_1}{R_3}$$

$$R_2 = \frac{R_1 R_4}{R_3}$$

$$K = \frac{l}{R_2} \times \frac{l}{A}$$

DPP-01

Ques. 5) $K = 6.3 \times 10^{-2} \Omega^{-1} \text{cm}^{-1}$
 $M = 0.1 \text{ M}$

$$\Lambda_m = \frac{K \times 1000}{M}$$

$$\Lambda_m = \frac{6.3 \times 10^{-2} \times 1000 \times 10}{0.1}$$

$$\Lambda_m = 630 \text{ ohm}^{-1} \text{cm}^2 \text{mol}^{-1}$$

Ques. 10)

$$\frac{l}{A} = 0.5 \quad R = 50 \quad N = 1$$

$$\Lambda_{eq} = \frac{K \times 1000}{N}$$

$$K = \frac{1}{R} \times \frac{l}{A}$$

$$K = \frac{1}{50} \times 0.5 = \frac{1}{100}$$

$$\Lambda_{eq} = \frac{\frac{1}{100} \times 1000}{1} = 10 \Omega^{-1} \text{cm}^2$$



Qw.11

$$\text{cell constant} = \frac{l}{A} = \frac{3}{4} \text{ cm}^{-1}$$

$$l = 3 \text{ cm}$$
$$A = 4 \text{ cm}^2$$

Qw.12

$$\frac{l}{A} = ? = G^*$$

$$\frac{K}{G} = G^*$$

$$K = 0.2 \text{ } \Omega^{-1} \text{ cm}^{-1}$$

$$\frac{0.2}{0.04} = \frac{20}{4} = 5 \text{ cm}^{-1}$$

$$G = 0.04 \text{ } \Omega^{-1}$$

$$K = G \times G^*$$

Qw.13

$$G = 1$$

$$K = 1$$

$$\frac{K}{G} = G^* = 1 \text{ cm}^{-1}$$

26/6/25

* Difference Between Metal Non metal and Metalloids

→ Metals

Metal have a valence shell of electron and metal are good conductor of heat electricity both.

Some non metal also behave like metal like- carbon graphite, carbon polymer, carbon black.

Some substance like glass ceramic have very poor conductivity, they are called as insulators.

Some substance like silicon, doping (doped silicon), Germanium, Arsenic they are called metalloids

They can act as conductor in certain condition

Some material have zero resistance and infinite conductance or conductivity they are called super conductor. At very low temperature 0°K to 15°K act as semi conductor. But Now a day's some ceramic oxide and mixed oxide are known to super conductor at 150°K They all are conductor of electricity or bad conductor electricity. They all need a piece of wire to flow electron (e^-).

This type of conductivity depends upon nature and structure of metal.

→ The Number of Valence electron per atom

→ Temperature:- When temperature increase resistance also increases but conductivity decreases.

$$T \propto R, T \propto \frac{1}{K}$$

* Ionic Conductivity

Like Metal conductivity one different type of conductivity which is depend upon concentration i.e. prepare a solution and conductance due to flow of ion. It is only present when it dissolved in polar solvent like water. Some of the solution doesn't dissolve or ionized in water. They do not show ionic conductance

→ It Depend upon :-

→ Nature of Electrolyte

→ Size of ions and there solvation

→ The Nature of Solvent and its viscosity

→ Concentration of the electrolyte

→ Temperature:- When temperature increase, ionic conductance also increases.

$$T \propto G(\text{ionic})$$

Ionic conductance

Weak Electrolyte

For strong Electrolytes

→ It is depend upon molar conductivity and dilution [Addition of H_2O].

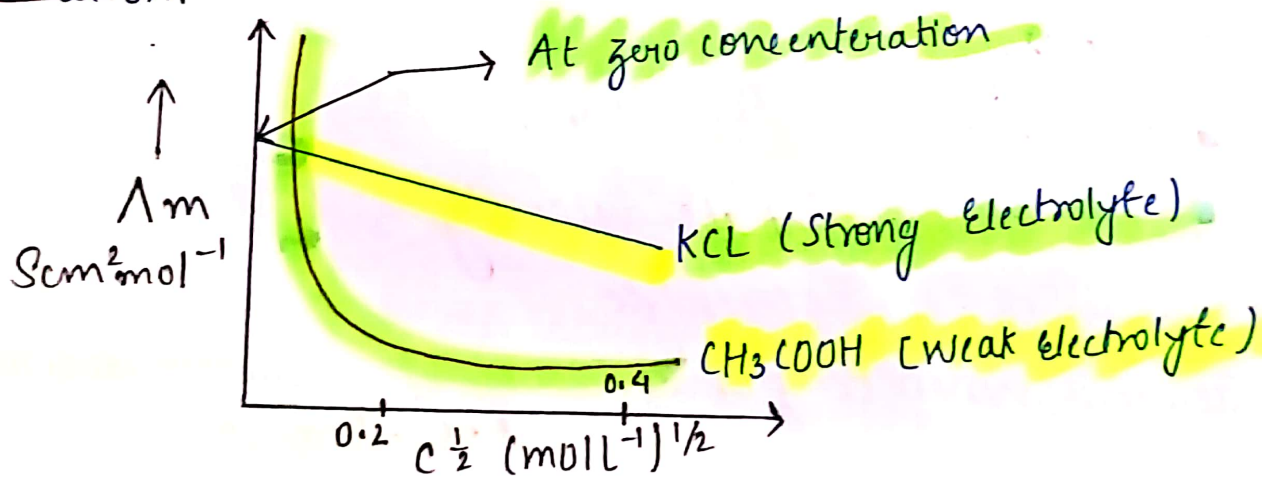
→ When adding of water increases conductivity increases for strong electrolyte upto ~~limited~~ limited. But for weak electrolyte it increases infinite. In that case we introduce a new a molar conductivity at infinite dilution denoted by $[\Lambda^{\infty}]$ or $[\Lambda_m^{\infty}]$

For Strong electrolyte They are represented as given

Equation Debye - Huckel - Onsager.

$$\left\{ \Lambda_m = \Lambda_m^0 - A C^{1/2} \right\}$$

If we carefully observe that graph it is clear shown that Weak electrolyte are less molar conductance than strong. If we add water [Dilution] conductivity increases for both. But after a some or certain time it observe that weak electrolyte molar conductance increase exponential and molar conductance goes to infinite that is known as Limiting Molar conductivity or conductivity at infinite dilution.



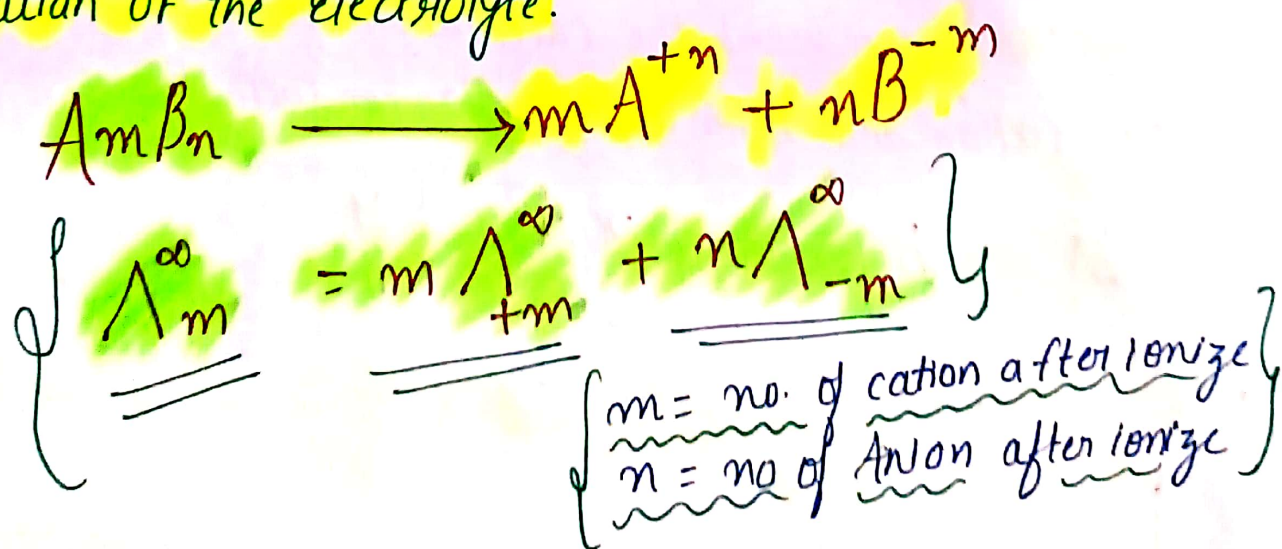
KOHRAUSCH'S LAW OF INDEPENDENT MIGRATION

According to Kohlrausch at infinite dilution when dissociation of electrolyte each ion make a definite contribution towards equivalent conductivity of any electrolyte irrespective of nature of ion associated with it

$$\Lambda_{eq}^{\infty} = \Lambda_{cation}^{\infty} + \Lambda_{anion}^{\infty}$$

According to NCERT, most imp

"The limiting molar conductivity of an electrolyte can be represented as the sum of individual contribution of the anion and cation of the electrolyte."



Degree of Ionization

Imp

$$\alpha = \frac{\Lambda_m}{\Lambda_m^\infty}$$

[Application of Kohlrausch's)

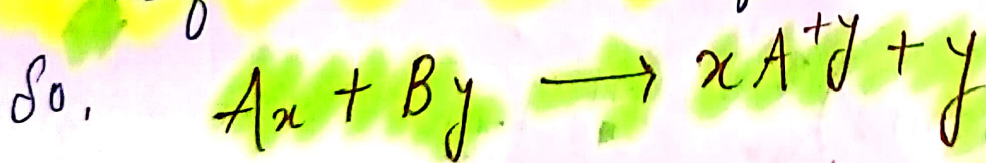
Ionization constant

$$K = \frac{C \alpha^2}{(1-\alpha)} = \frac{C \Lambda_m^2}{\Lambda_m^{\infty 2} \left(1 - \frac{\Lambda_m}{\Lambda_m^\infty}\right)} = \left[\frac{C \Lambda_m^2}{\Lambda_m^\infty (\Lambda_m^\infty - \Lambda_m)} \right]$$

Imp

For weak Electrolytes They don't ionise 100%. So we can calculate Degree of Ionization by calculating molar conductivity and molar conductivity at infinite dilution.

For weak Electrolyte we also calculate ionization constant that is denoted by $[K]$. If we want the calculation of Equivalent conductivity in terms of cation and anion of molar conductivity.



Imp $\left\{ \Lambda_{eq}^0 = \frac{1}{x} \Lambda_m^{\infty} \text{ cation} + \frac{1}{y} \Lambda_m^{\infty} \text{ anion} \right\}$

Factor Affecting Conductivity

- It doesn't depend on the concentration (κ)
- It only depends upon resistance, temperature and interionic attraction
- For Strong Electrolyte Conductivity decreases on dilution
- It also depends upon the number of ions which is dissociated in the solution

Factor affecting Electrolyte Conductance $[\Lambda]$

- Inter ionic attraction
The inter ionic attraction of solution is more then the conductance will less.
- Polarity of solution
Solution has High dielectric constant then the ionization and conductance will be higher.
- Viscosity of medium
On Increasing the viscosity of medium the conductance decreases
- Temperature
Temperature can done two effect
 - ① When Temperature increases the K.E of molecules increase flow of ions also increase
 - ② When Temperature decrease Viscosity

→ Hydrate size

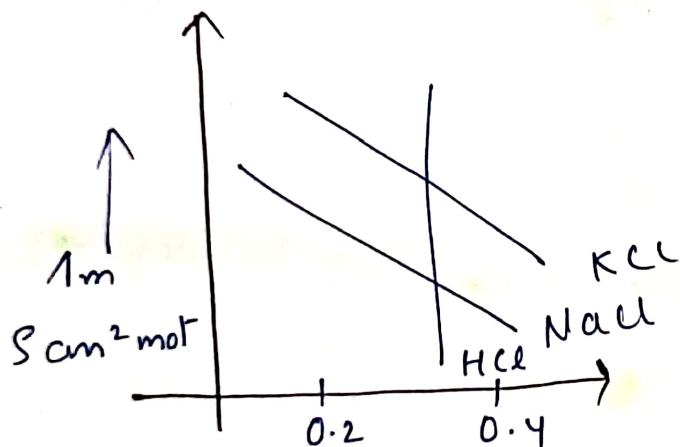
When size of ions increases conductivity decreases due to the heaviness of ions.

→ Dilution

① On Increasing Dilution conductance [G] increases

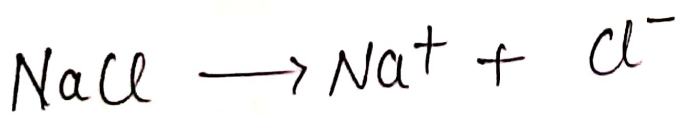
② Molar conductivity increase for weak electrolyte

[α] increases, specific conductance decreases and molar conductivity and equivalent conductivity also increases. Due to number of ions per unit area decreases then the speed of ions increases.

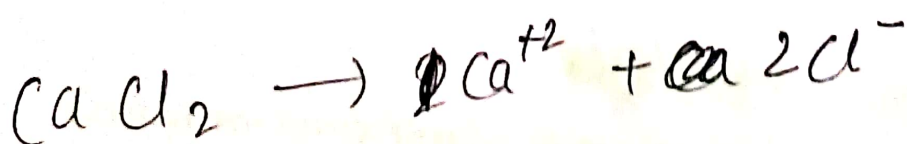


$$\Lambda_m \propto C^{1/2}$$

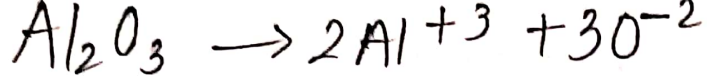
$$\Lambda_m = \Lambda_m^\infty - AC^{1/2}$$



$$\Lambda_m^\infty = 1 \times \Lambda_{\text{Na}^+}^\infty + 1 \times \Lambda_{\text{Cl}^-}^\infty$$



$$\Lambda_m^\infty = 1 \times \Lambda_{\text{Ca}^{+2}}^\infty + 2 \times \Lambda_{\text{Cl}^-}^\infty$$



$$\Lambda_m^\infty = 2 \times \Lambda_{Al^{+3}}^\infty + 3 \times \Lambda_{O^{-2}}^\infty$$

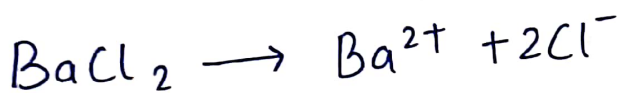
DPP-03

Pg-21

Q.1

$$\Lambda_{Ba^{+2}}^\infty = 127 \Omega^{-1} cm^2 g eq^{-1}$$

$$\Lambda_{Cl^{-}}^\infty = 76 \Omega^{-1} cm^2 g eq^{-1}$$



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

$$\Lambda_{BaCl_2}^\infty = 127 + 76$$

$$\Lambda_{BaCl_2}^\infty = 203 \Omega^{-1} cm^2 g eq^{-1}$$

Ques 2

$$\alpha = \frac{\Lambda_m}{\Lambda_{eq}^{\infty}} \rightarrow \frac{K \times 1000}{M}$$

$$\rightarrow \Lambda_{CH_3COO^-}^{\infty} + \Lambda_{H^+}^{\infty}$$

$$\Lambda_{CH_3COOH}^{\infty} = \Lambda_{CH_3COO^-}^{\infty} + \Lambda_{H^+}^{\infty}$$

$$= 42 + 288.42$$

$$\Lambda_{CH_3COOH}^{\infty} = 320.42 \text{ } \Omega^{-1} \text{cm}^2 \text{mol}^{-1}$$

$$\Lambda_m = 12.8 \text{ } \Omega^{-1} \text{cm}^2 \text{mol}^{-1}$$

$$\alpha = \frac{12.8}{320.42} \frac{\Omega^{-1} \text{cm}^2 \text{mol}^{-1}}{\Omega^{-1} \text{cm}^2 \text{mol}^{-1}} = 0.0387$$

$$\alpha = 0.0387 \times 100\% = 3.87\%$$

$$\approx 3.9\%$$

Ques 3

$$\Lambda_{NaCl}^{\infty} = \Lambda_{Na^+}^{\infty} + \Lambda_{Cl^-}^{\infty} = 126.45 \text{ } \Omega^{-1} \text{cm}^2 \text{mol}^{-1}$$

$$\Lambda_{HCl}^{\infty} = \Lambda_{H^+}^{\infty} + \Lambda_{Cl^-}^{\infty} = 426.16 \text{ } \Omega^{-1} \text{cm}^2 \text{mol}^{-1}$$

$$\Lambda_{CH_3COONa}^{\infty} = \Lambda_{CH_3COO^-}^{\infty} + \Lambda_{Na^+}^{\infty} = 91 \text{ } \Omega^{-1} \text{cm}^2 \text{mol}^{-1}$$

$$\Lambda_{CH_3COOH}^{\infty} = \Lambda_{CH_3COO^-}^{\infty} + \Lambda_{H^+}^{\infty}$$

$$\Lambda_{\text{CH}_3\text{COONa}}^\infty \neq \Lambda_{\text{HCl}}^\infty - \Lambda_{\text{NaCl}}^\infty = \Lambda_{\text{CH}_3\text{COOH}}^\infty$$

$$91 + 426.16 - 125.45 = \Lambda_{\text{CH}_3\text{COOH}}^\infty$$

$$\Lambda_{\text{CH}_3\text{COOH}}^\infty = 390.71 \text{ } \Omega^{-1} \text{cm}^2 \text{mol}^{-1}$$

Q.4)

$$\Lambda_{\text{ClCH}_2\text{COONa}}^\infty = \Lambda_{\text{ClCH}_2\text{COO}^-}^\infty + \Lambda_{\text{Na}^+}^\infty = 224 \text{ } \Omega^{-1} \text{cm}^2 \text{g eq}^{-1}$$

$$\Lambda_{\text{NaCl}}^\infty = \Lambda_{\text{Na}^+}^\infty + \Lambda_{\text{Cl}^-}^\infty = 38.7 \text{ } \Omega^{-1} \text{cm}^2 \text{g eq}^{-1}$$

$$\Lambda_{\text{HCl}}^\infty = \Lambda_{\text{H}^+}^\infty + \Lambda_{\text{Cl}^-}^\infty = 203 \text{ } \Omega^{-1} \text{cm}^2 \text{g eq}^{-1}$$

$$\Lambda_{\text{ClCH}_2\text{COO}^-}^\infty + \Lambda_{\text{H}^+}^\infty$$

$$\Lambda_{\text{ClCH}_2\text{COONa}}^\infty \neq \Lambda_{\text{HCl}}^\infty \quad \Lambda_{\text{NaCl}}^\infty = \Lambda_{\text{ClCH}_2\text{COOH}}^\infty$$

$$224 + 203 - 38.7 = \Lambda_{\text{ClCH}_2\text{COOH}}^\infty$$

$$\Lambda_{\text{ClCH}_2\text{COOH}}^\infty = 388.5 \text{ } \Omega^{-1} \text{cm}^2 \text{g eq}^{-1}$$

Q.5)

$$\Lambda_{\text{NaBr}}^\infty = \Lambda_{\text{Na}^+}^\infty + \Lambda_{\text{Br}^-}^\infty = ?$$

$$\Lambda_{\text{NaCl}}^\infty = \Lambda_{\text{Na}^+}^\infty + \Lambda_{\text{Cl}^-}^\infty = 126 \text{ } \Omega \text{cm}^2 \text{mol}^{-1}$$

$$\Lambda_{\text{KBr}}^\infty = \Lambda_{\text{K}^+}^\infty + \Lambda_{\text{Br}^-}^\infty = 152 \text{ } \Omega \text{cm}^2 \text{mol}^{-1}$$

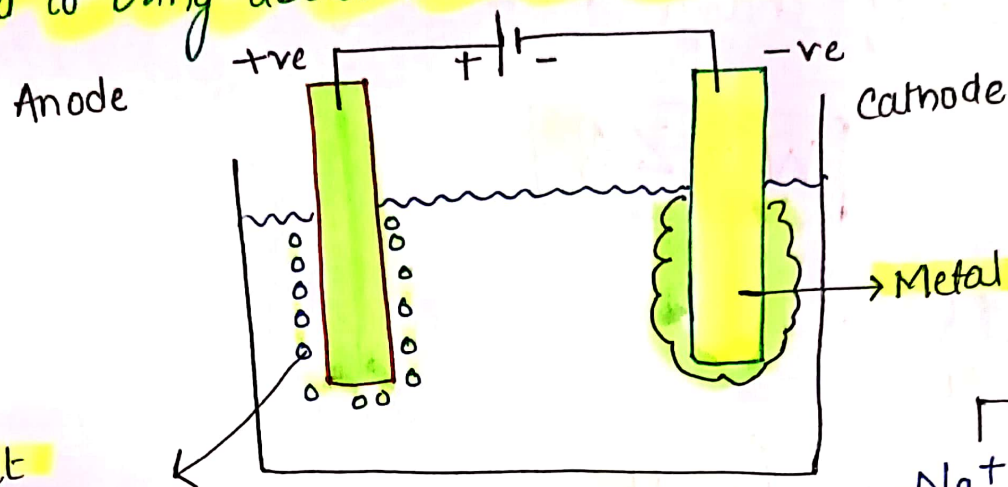
$$\Lambda_{\text{KCl}}^\infty = \Lambda_{\text{K}^+}^\infty + \Lambda_{\text{Cl}^-}^\infty = 150 \text{ } \Omega \text{cm}^2 \text{mol}^{-1}$$

$$\Lambda_{\text{NaCl}}^\infty + \Lambda_{\text{KBr}}^\infty - \Lambda_{\text{KCl}}^\infty = \Lambda_{\text{NaBr}}^\infty$$

$$\Lambda_{\text{NaBr}}^\infty = 126 + 152 - 150 = 128 \text{ } \Omega \text{cm}^2 \text{mol}^{-1}$$

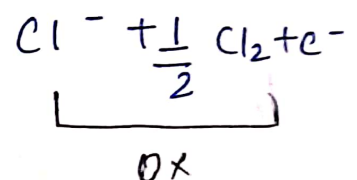
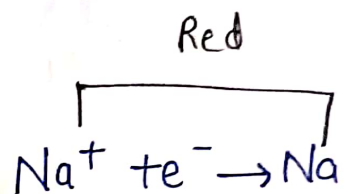
* ELECTROLYTIC CELL

"In an Electrolytic cell external source of voltage is used to bring about a chemical reaction".



Metal $\rightarrow$ Cation
 Na^+

Non-Metal $\rightarrow$ Anion
 Cl^-



Michael Faraday, He was the first Scientist who discovered quantitative aspect of electrolysis.

$\rightarrow$ Faraday's first law

"The amount of Chemical Reaction which occurs at electrode during electrolysis by a current is proportional to the quantitative of Electricity passed through the Electrolytes."

Largest Unit of Charge = Faraday $1F$
= 1 mole of e^- flow

$$= 6.02 \times 10^{23} \times 1.6 \times 10^{-19} \text{ C}$$

$$1F = 96480 \text{ coulombs}$$

$$1F \approx 96500 \text{ coulombs}$$

$$W \propto Q$$

W = amount of metal deposited at cathode

Q = amount of charge flow through electrode

$$W = ZQ$$

$$\therefore \frac{dQ}{dt} = I$$

$$\int dQ = \int I \cdot dt$$

$$Q = It$$

Z = Equivalent constant

$$Z = \frac{E}{F} \rightarrow \text{Equivalent weight}$$

$$W = ZIT$$

$$W = ZQ$$

$$\because Q = ne$$

$$W = Zne$$

$$W = \frac{EIt}{F}$$

$$W = \frac{EIt}{96500}$$

$$E = \frac{\text{atomic mass}}{\text{Valency Factor} \times [\text{no. of } e^- \text{ exchanges}]}$$

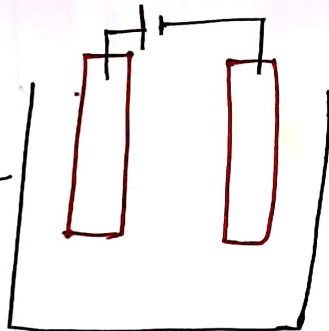
$$W = \frac{\text{at. wt} \times I \times t}{V \cdot F \times 96500}$$

Faraday First Law

→ Faraday Second Law

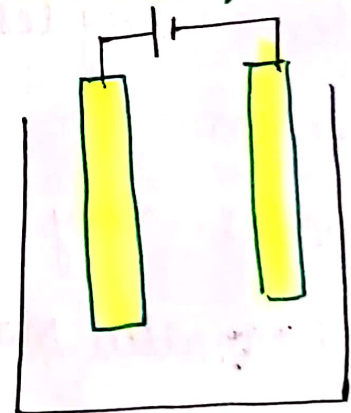
"The amount of different substances liberated by same amount of electricity passing through the electrolytic solution are proportional to their chemical equivalent weight".

$$W_1 = \frac{E_1 \times I \times t}{F}$$



I same
t = same

$$W_2 = \frac{E_2 It}{F}$$



I same
t same

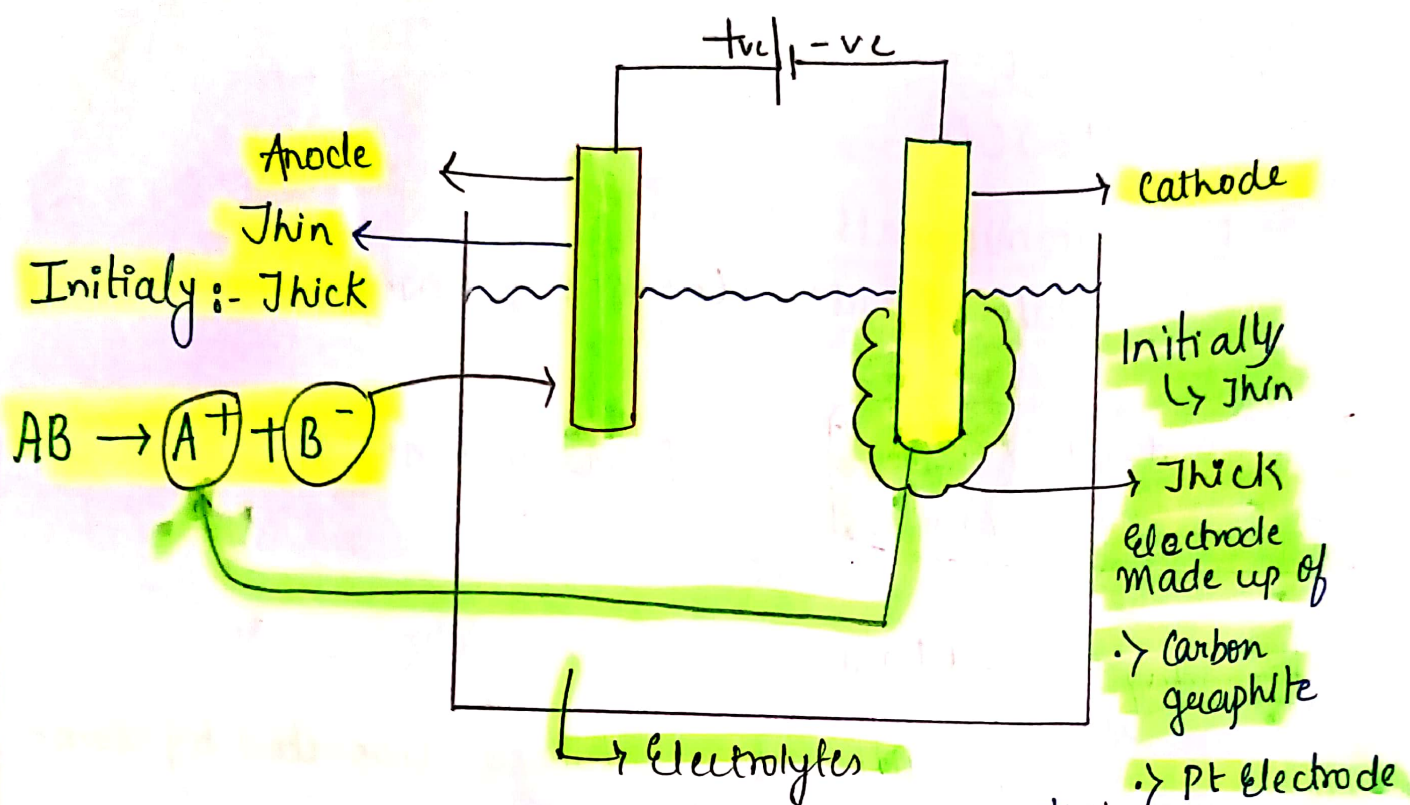
$$W \propto E$$

$$W_1 = W_2$$

$$\frac{W_1}{W_2} = \frac{E_1}{E_2}$$

1/07/25

PRODUCT OF ELECTROLYSIS



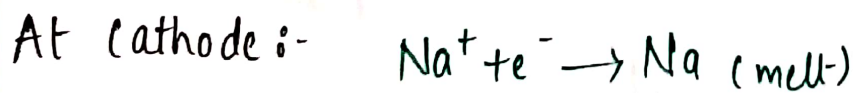
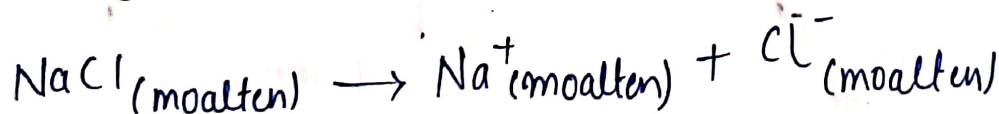
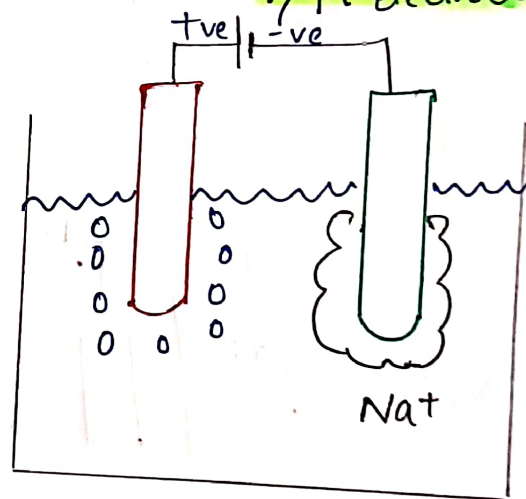
Oxidation

Anode :- Very less E° value.

Reduction

Cathode :- Very high value of E°

→ (i) Molten NaCl

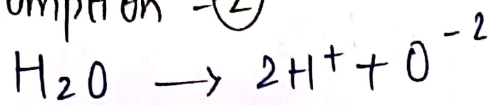


Aqueous solution of NaCl

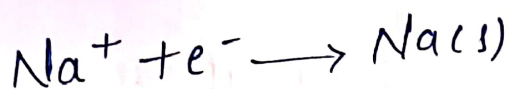
→ Assumption - (1)



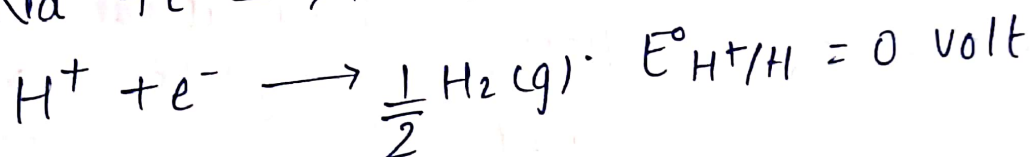
→ Assumption - (2)



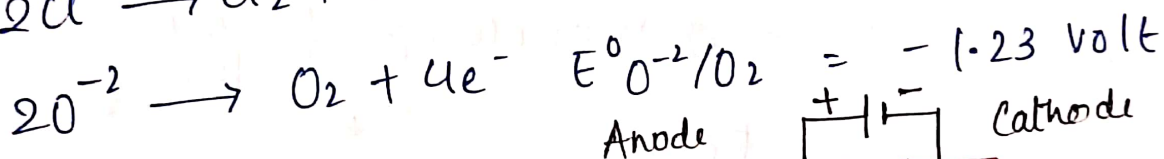
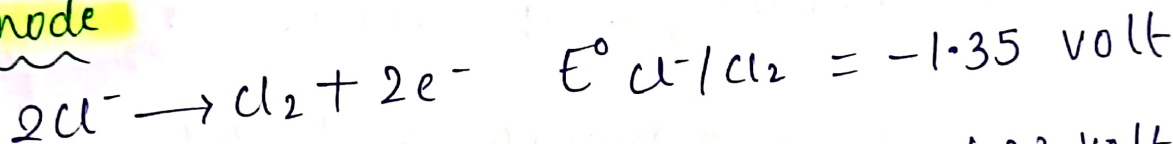
At cathode E°_{max}



$$E^\circ_{\text{Na}^+/\text{Na}} = -2.71 \text{ Volt}$$

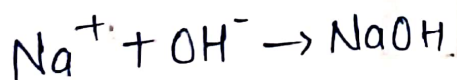


At anode

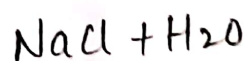
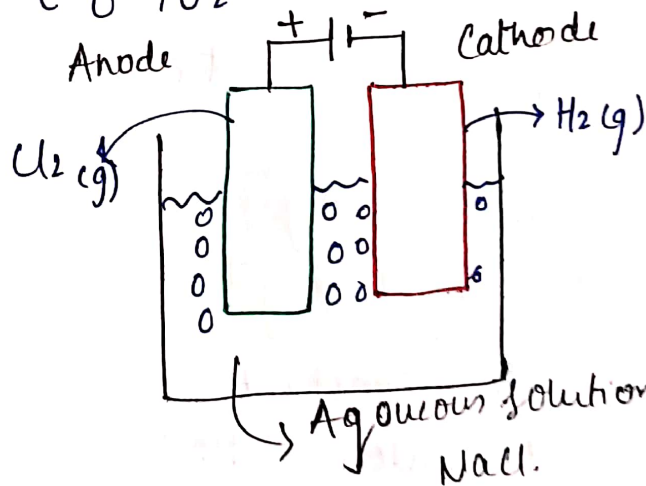


at cathode - H_2

at Anode - Cl_2



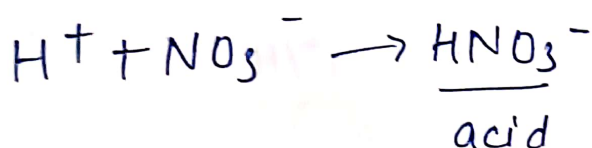
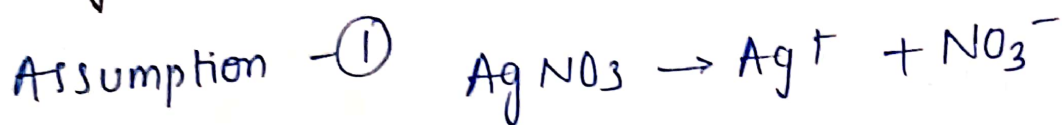
pH of the solution increases



$$P_g = 60 \quad [\text{NCERT}]$$

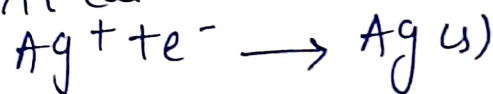
Q.2.18

1) AgNO_3 with Ag electrode

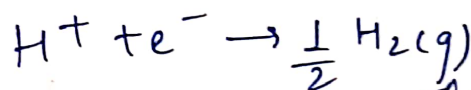


pH = decreases

At Cathode

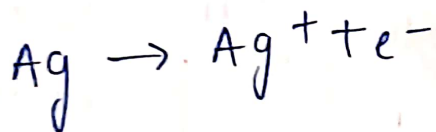


$$E^\circ = +0.79 \text{ V}$$



$$E^\circ = 0 \text{ V}$$

At Anode :- ROD is made up of Ag

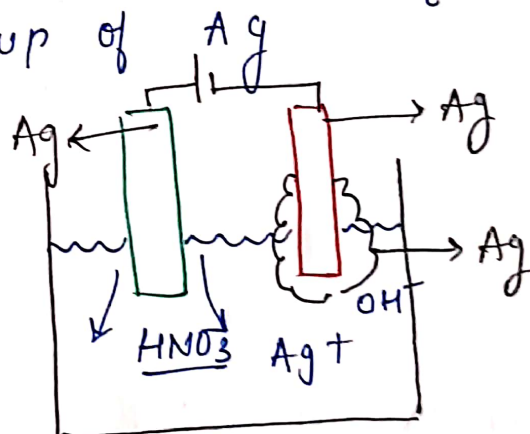


Cathode - Ag

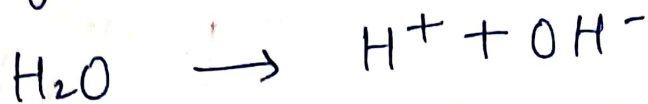
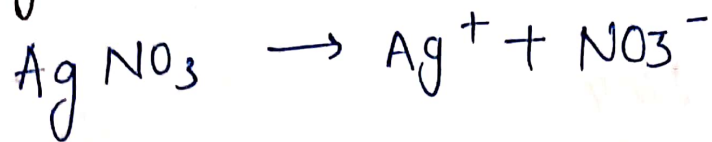
Anode - Ag^+

pH = constant

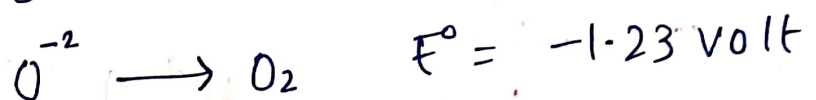
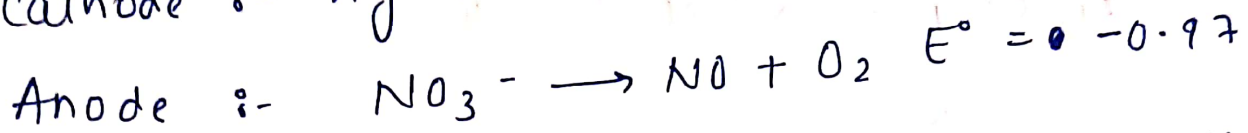
Formation of HNO_3



(II) AgNO_3 with Pt electrode

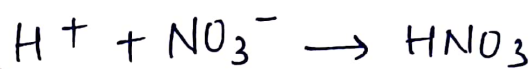


Cathode :- Ag its E° are greater than H_2

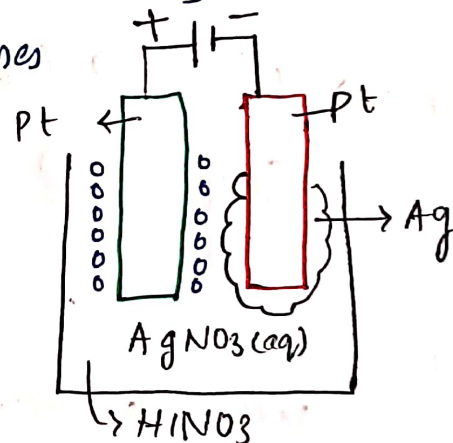


Anode - O_2

Cathode - Ag

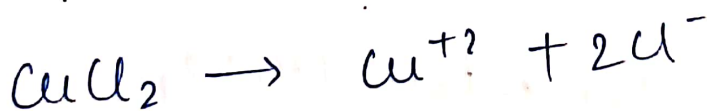


pH decreases



(IV) An aqueous solution of CuCl_2 with Pt electrode

Assn ①



Assn ②



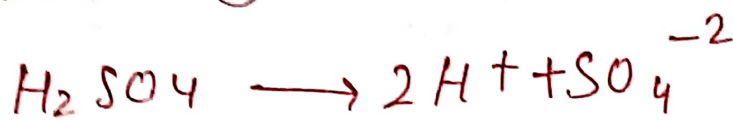
Cathode - Cu(s)

Anode - Cl(g)

Q.1.28

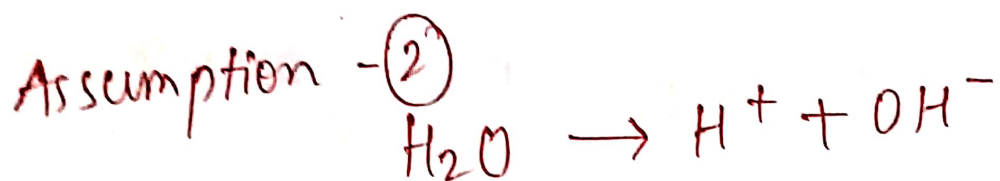
(11) Dilute H_2SO_4

Assumption - (1)

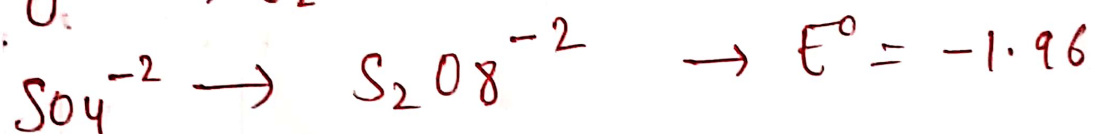
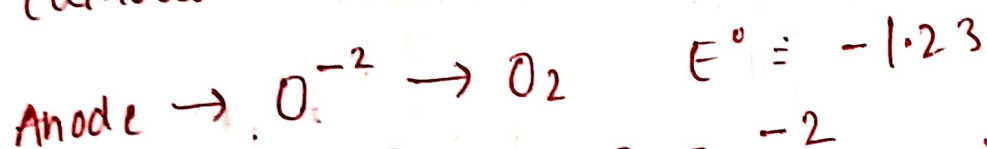


(very very less)

Assumption - (2)



Cathode - H_2



Note

50 Stars

When we take dilute H_2SO_4 cathode form H^+ and anode form O_2 51 Stars

When we take concentrated H_2SO_4

Cathode - H_2

Anode - $S_2O_8^{-2}$

Electrolyte (electrode)	Cathode	Anode	pH
Molten NaCl (pt)	Na	Cl_2	Same
Aq. NaCl (pt)	H_2	Cl_2	pH increase NaOH
AgNO_3 [Ag]	Ag	Ag^+	pH same
AgNO_3 [pt]	Ag	O_2	pH decreases
Dil. H_2SO_4 (pt)	H_2	O_2	pH same.
Conc. H_2SO_4 (pt)	H_2	$\text{S}_2\text{O}_8^{-2}$	pH decrease [$\text{H}_2\text{S}_2\text{O}_8$] [per oxo sulphuric acid]
CuCl_2 [pt]	Cu	Cl_2	pH same

3/07/2025

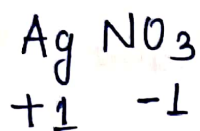
DPP-10

Pg = 41

Ques 1)

Weight of substance deposited at cathode = 108 g

Atomic weight of substance = 108 g



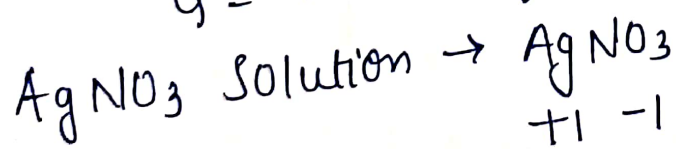
Valency of metal = Ag = 1

$$W = \frac{\text{At. wt} \times Q}{\text{Valency} \times F}$$

$$108 = \frac{108 \times Q}{1 \times F}$$

$$Q = 1 \text{ Faraday}$$

Ques 2) Q = Amount of electron = 9.65 coulomb



Valency = 1, At wt of metal = Ag = 108 g

$$W = \frac{\text{At wt} \times Q}{\text{Val} \times F}$$

$$W = \frac{108 \times 965}{1 \times 96500} = \frac{108}{10} = 10.8 \text{ mg}$$

Ques. 3) atomic weight of Al = 27 g

Q = Amount of current = 5F

$$\text{Valency} = Al^{+3} = 3$$

$$W = \frac{\text{At wt} \times Q}{\text{Val} \times F}$$

$$W = \frac{27 \times 5F}{3 \times F}$$

W = 45 g of Aluminium.

Ques. 4)

$$Z = \frac{E}{F} = \frac{\text{At wt (g)}}{\text{Val (unitless)} F [\text{Columb}]}$$

$$Z = \frac{g}{\text{columb}}$$

Ques. 5) Faraday = 1 mole of e^-

$$\frac{e}{\text{mole}} = \frac{\text{Columb}}{\text{mole}}$$

Ques. 6)

$$(3) \boxed{m = ect}$$

Ques. 7) current passes them $I = 5(A)$

Time required (t) = 40 min

$$= 40 \times 60 = 2400 \text{ sec.}$$

Solution taken = $ZnSO_4$
+2 -2

Valency of $ZnSO_4$ = $Zn = 2$

$$W = \frac{\text{At} \times I \times t}{\text{Val} \times 96500}$$

$$W = \frac{65 \times 5 \times 2400}{2 \times 96500}$$

$$\boxed{W = 4.065 \text{ gm}}$$



Qw. 8)

25 Stars

$$W \propto I$$

$$W \propto t$$

$$\frac{W_1}{W_2} = \frac{I_1 \times t_1}{I_2 \times t_2}$$

$$\frac{m}{W_2} = \frac{4 \times 2 \times 60}{6 \times 40}$$

$$\boxed{W_2 = \frac{m}{2}}$$

Qw. 9) Amount of substance deposited at cathode = 1.8 g

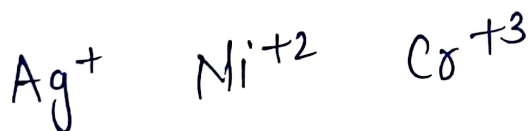
current passing = 3 Amp

time duration = $50 \times 60 = 3000$ sec

$$W = \frac{E \times I \times t}{F} = \frac{1.8 \times 96500}{3 \times 3000} = E$$

$$\boxed{E = 19.3 \text{ g eq}^{-1}}$$

Qw. 10)

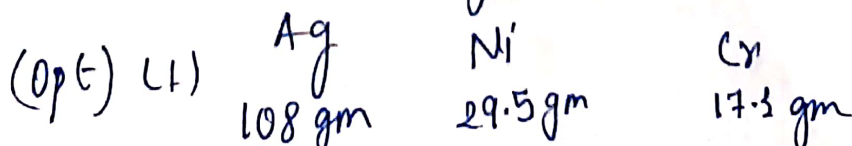


$$\frac{\text{At. wt}}{\text{valency}} = \frac{108}{1} \quad \frac{59}{2} \quad \frac{52}{3}$$

$$= 108 \text{ g} = 29.5 \text{ g} = 17.33 \text{ g}$$

$$W = \frac{E \times I \times t}{F}$$

$$W = E = \frac{\text{At. wt}}{\text{valency}}$$



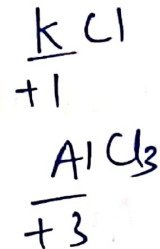
Ques. 11) (4) None of These.

Ques. 12) 2nd Law

I same, t same

amount of Charge Current same

$$\frac{w_1}{w_2} = \frac{At_1 \times Val_2}{At_2 \times Val_1}$$



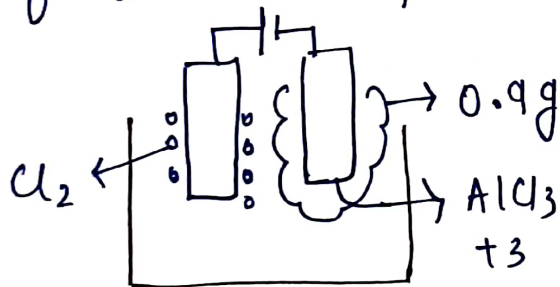
$$\frac{19.5}{w_2} = \frac{39 \times 3}{27 \times 7}$$

$$w_2 = \frac{19.5 \times 27}{39 \times 3}$$

$$w_2 = 4.5g$$

mass of aluminum deposited at cathode $w_2 = 4.5g$

Ques. 13)



t = same I = same

$$\frac{w_1}{w_2} = \frac{At_1 \times Val_2}{At_2 \times Val_1}$$

$$At wt_1 = Al = 27$$

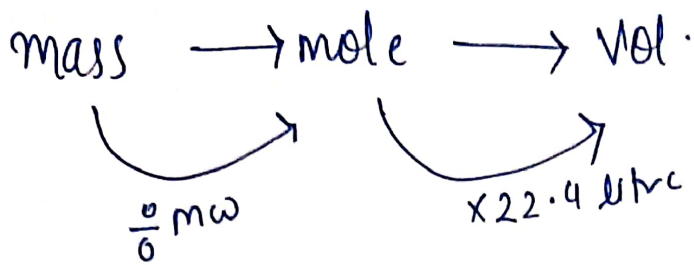
$$At wt_2 = Cl_2 = 71$$



$$\frac{0.9}{w_2} = \frac{27 \times 2}{71 \times 3}$$

$$w_2 = \frac{71 \times 3 \times 0.9}{27 \times 2}$$

$$w_2 = 3.55g$$



$$\frac{3.55}{71} \times 22.4 = 1.12 \text{ litre}$$

Ques. 14) $1 \text{ gm atom} = 1 \text{ mole} = 27 \text{ g}$

$$w = \frac{At \times n \times e}{val \times e \times N_A}$$

$$27 = \frac{27 \times n}{3 \times N_A}$$

$$\boxed{n = 3 N_A}$$

$$N_A = 6.02 \times 10^{23}$$

Avogadro's Number

Ques. 15)

$$w = 0.5 E$$

$$w = \frac{E \times Q}{F}$$

$$0.5 E = E \times \frac{Q}{F}$$

$$Q = 0.5 F$$

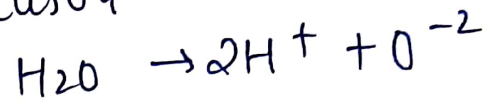
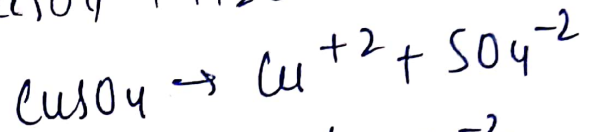
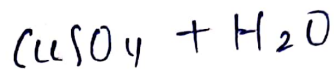
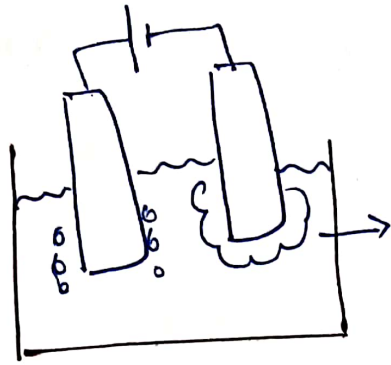
$$Q = 0.5 \times 96500$$

$$\boxed{Q = 48250 \text{ Coulomb}}$$

DPP - 11

Pg - 44

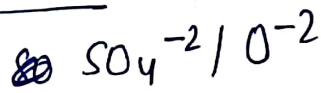
Qu. 1)



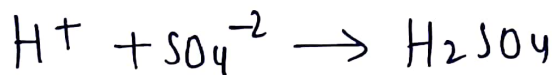
Cathode $\rightarrow$ Cation
 $\text{H}^+ / \text{Cu}^{+2}$

Cu :- deposited

Anode $\rightarrow$ Anion



SO_4^{-2} :- less / $\text{O}^{-2} \rightarrow$ more $\rightarrow \text{O}_2$



pH decreases

Qu. 2)

$$I = 2 \text{ ampere}$$

$$t = 5 \text{ hr} = 5 \times 3600 =$$

$$W = 22 \text{ g}$$

$$At = 177$$

$$Val = 2$$

$$W = \frac{At \times I \times t}{Val \times F}$$

$$22 = \frac{177 \times 2 \times 5 \times 3600}{Val \times 96500}$$

$$Val = 2.87 \approx +3$$

5/07/2025

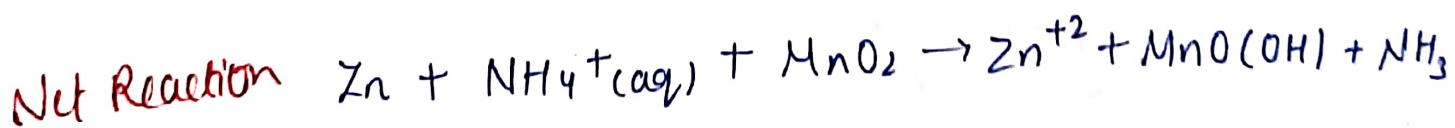
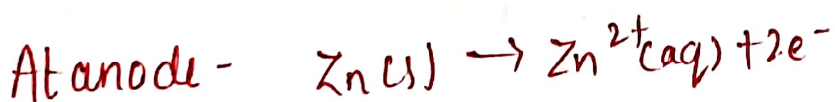
* Battery

Combination of galvanic cells in series are used as a source of electrical energy.

- (i) Primary batteries are non rechargeable batteries such as Leclanche cell and Dry cell.
- (ii) Secondary batteries are chargeable cells involving reversible reaction.
Example:- Lead Storage battery and Nickel-cadmium cells.

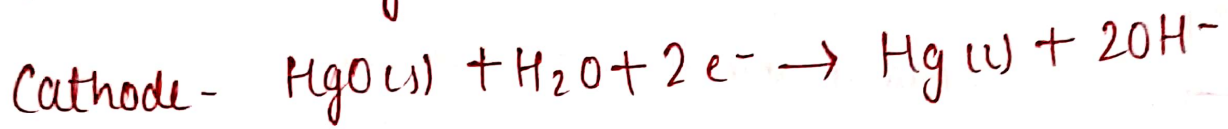
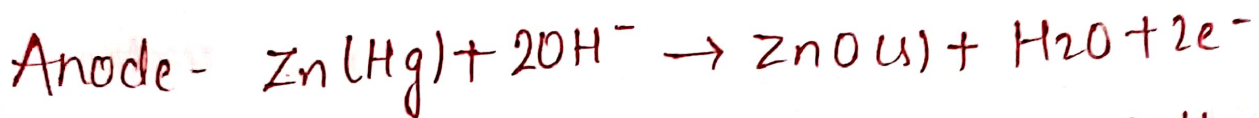
Dry cell [Leclanche cell]

The anode consists of a zinc container and the cathode is a graphite electrode surrounded by powdered MnO_2 and C. The space is filled with paste of NH_4Cl and ZnCl_2 .



Mercury Cell

It consists of zinc mercury amalgam as anode and a paste of HgO and Carbon as a cathode. The Electrolyte is a paste of KOH and ZnO. The Electrodes are:-



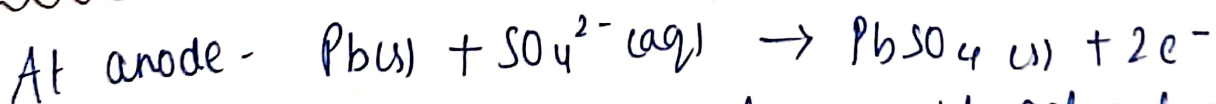
The cell potential is approximately 1.35 Volt and remains constant as the ionic concentration of the solution is not changed during its life.

Lead Storage Cell

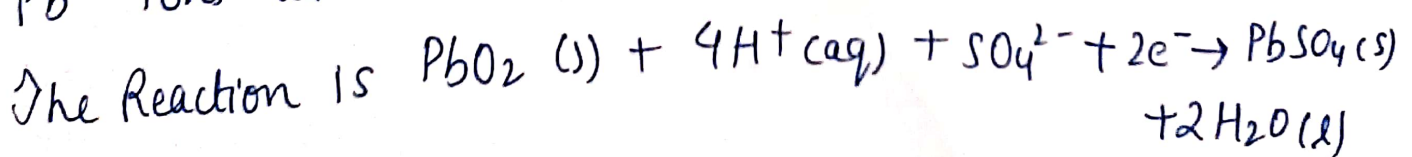
Anode - Spengy Lead Cathode - Lead packed with lead dioxide

Electrolyte - Aqueous solution of H_2SO_4 (38%)

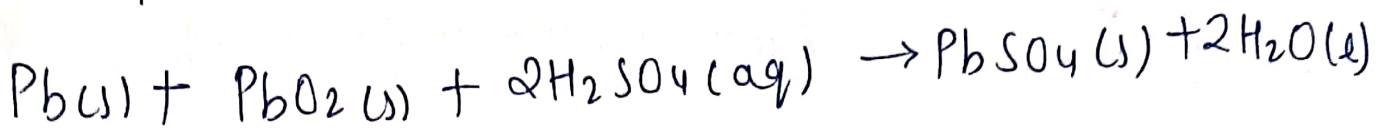
Discharge reaction of cell



At cathode - PbO_2 filled in lead grid get reduced to Pb^{2+} ions which combines with SO_4^{2-} ions to form $\text{PbSO}_4(\text{s})$.



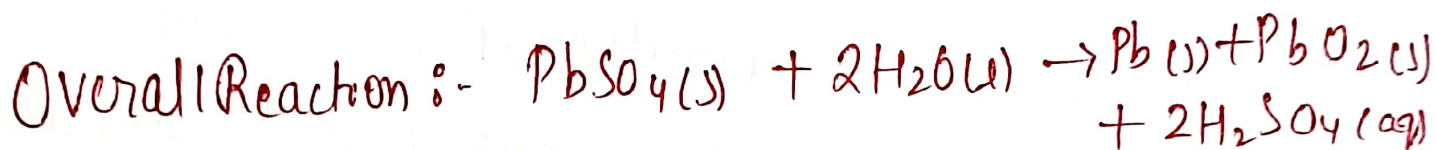
Complete Cell Reaction



Recharge reaction of cell (Charging)

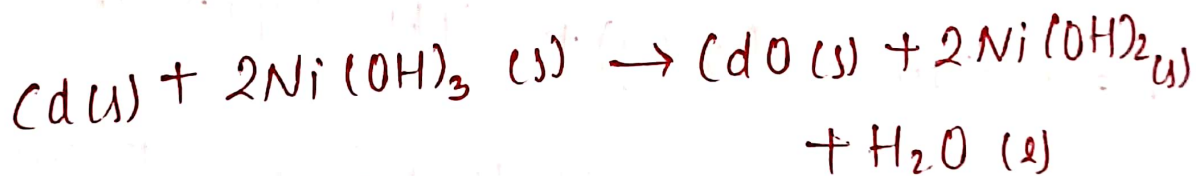
It changes the direction of electrode reaction.

PbSO_4 accumulated at cathode get reduced to Pb .



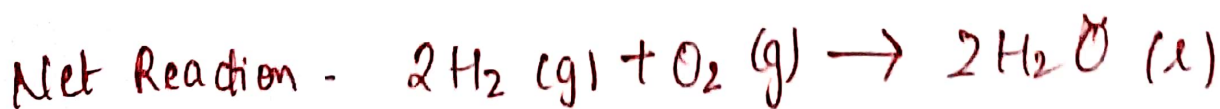
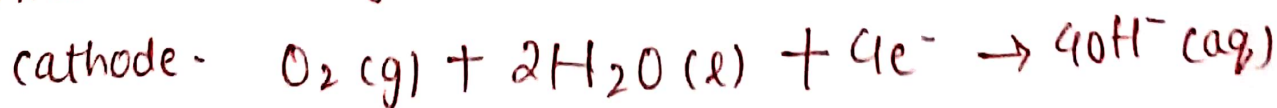
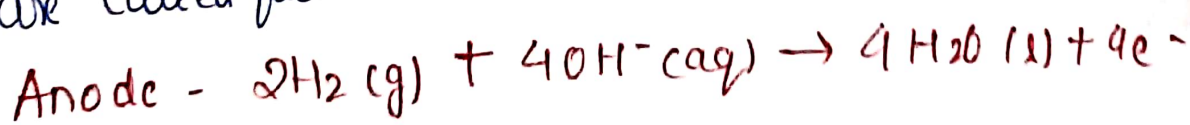
Nickel-Cadmium cell

Which has longer life than the lead storage cell but is costly. Here, the overall reaction during discharging.



Fuel cells

Electrical cells that are designated to convert the energy from the combustion of fuels such as hydrogen, carbon monoxide or methane directly into electrical energy are called fuel cells. The cell reactions are-



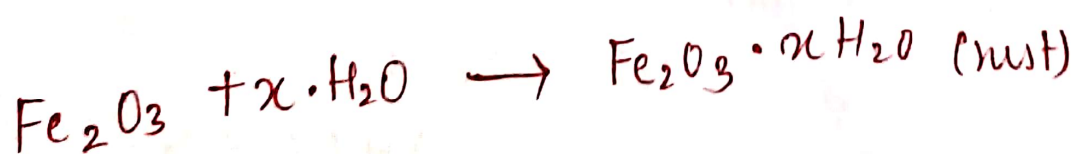
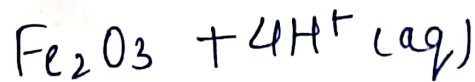
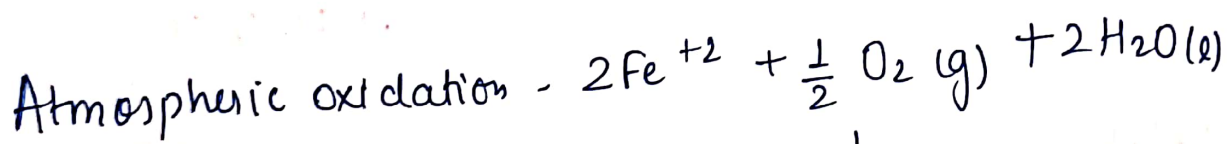
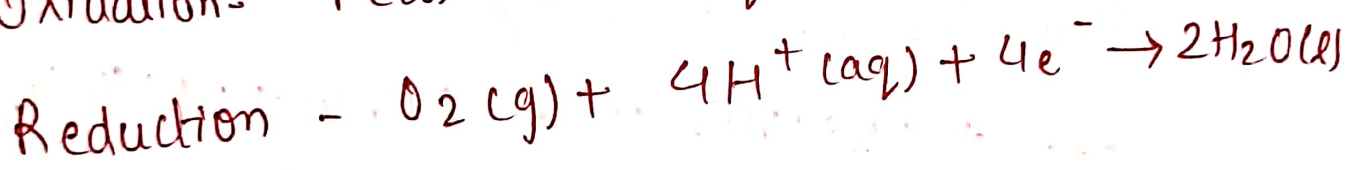
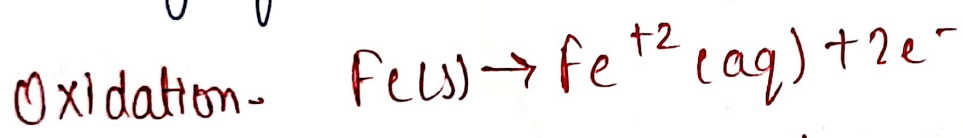
Advantages

High efficiency, free from pollution (eco-friendly), continuous source of energy.

Corrosion

The process of slow conversion of metals into their undesirable compounds (usually oxides) by reaction with moisture and other gases present in the atmosphere.

Rusting of iron



Prevention of Corrosion

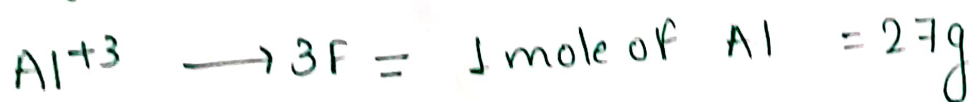
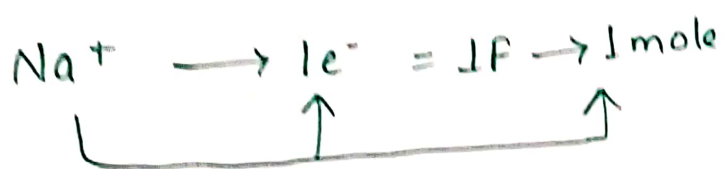
(1) Barrier protection:- By covering the surface with paint or a thin film of grease or by some chemicals (eg-bisphenol).

Abn. ~~==~~ concepty

1 mole of e^- = 1F

$$2 \text{ mole of } e^- = 2F$$

$$4 \text{ mole of } e^- = 4F$$



—पू कि

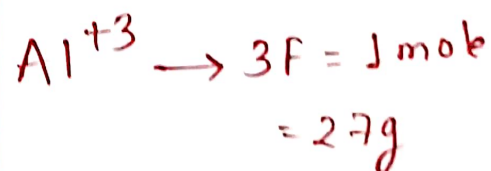
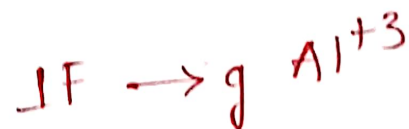
$$10 \text{ रु} = 3 \text{ केले}$$

10 ₹ = 3 केले

$$1 \text{ र} = \frac{3}{10} \text{ के ल}$$

$$50 \text{ ₹} = \frac{3}{10} \times 50$$

50 ₹ = 15 ~~के~~ का



$$3F = 27g$$

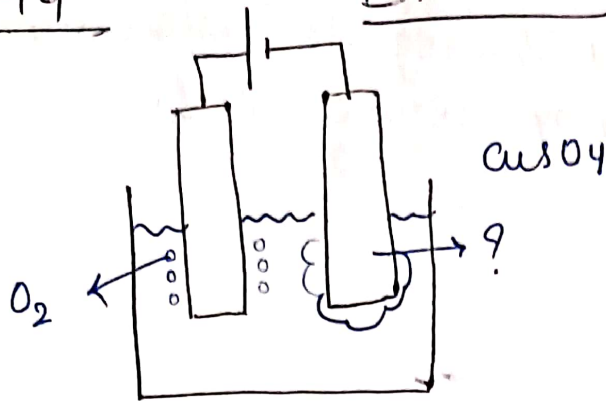
$$JF = \frac{27}{3} = 9g$$

6/07/2025

Pg-44

DPP-11

Q.4)



$$E_1 = \frac{32}{4} = 8$$

$$E_2 = \frac{63.5}{2}$$

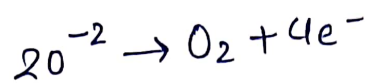
$O_2 = 11.2$ litres at STP

$$\frac{w_1}{w_2} = \frac{E_1}{E_2} \Rightarrow \frac{16}{w_2} = \frac{8 \times 2}{63.5}$$

$$w = 63.5 g$$

Vol $\rightarrow$ mole $\rightarrow$ mass
 $\div 22.4 \text{ litre}$ $\times \text{mw}$

$$\frac{11.2}{22.4} \times 32 = 16 g$$



Q.5)

Q.6)

$H_2 \rightarrow \text{Vol} = 112 \text{ ml}$

Vol $\rightarrow$ mol $\rightarrow$ weight
 $\div 22.4 (l)$ $\times \text{mw}$

$$\frac{112}{22400} = \frac{1}{200} \text{ mole} \times 2 = \frac{1}{100} g = 0.01 g$$

$$w = \frac{At \cdot w \times I \times t}{\text{Vol} \times F}$$

$$0.01 = \frac{2 \times I \times 965}{2 \times 96500}$$

$$I = 1 \text{ Amp}$$

Ques. 7) $V = 200 \text{ ml}$

$$N = 0.1$$

AgNO_3 $I = 0.1 \text{ Amp}$

$n_{\text{factor}} = 1$ $A_g = 108 \text{ g}$

time

$$N = \frac{w \times n_{\text{factor}} \times 1000}{M_w \times V}$$

$$0.1 = \frac{w \times 1 \times 1000}{108 \times 200}$$

$$w = \frac{10.8 \times 2}{10} = 2.16 \text{ g}$$

$$\text{half of silver} = \frac{2.16}{2} = 1.08$$

$$w = \frac{At \times I \times t}{\text{Val} \times F}$$

$$1.08 = \frac{108 \times 0.1 \times t}{1 \times 96500}$$

$$t = 9650 \text{ sec}$$

Ques. 8 112 ml H_2 $1 \text{ ml} = 1 \text{ cm}^3$

Vol $\rightarrow$ mole $\rightarrow$ mass
 $\div 22.4$ $\times M_w$

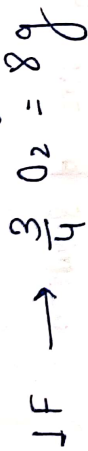
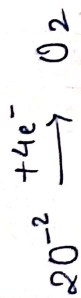
$$\frac{112}{22400} = \frac{1}{200} \times 2 = 10^{-2}$$

$$Q = 9650 \text{ C}$$

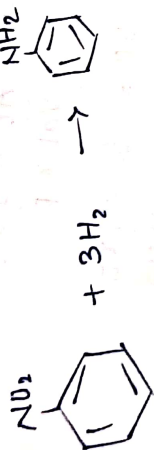
$$w = \frac{At \times Q}{\text{Val} \times F} = 0.01 = \frac{2 \times Q}{2 \times 96500}$$



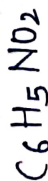
Qu. 9



Qu. 10



GF = 1 mole Nitro benzene



$$6 \times 12 = 72$$

$$5 \times 1 = 5$$

$$14 \times 1 = 14$$

$$16 \times 2 = 32$$

$$\frac{123g}{123g}$$

$$\text{mole of Nitrobenzene} = \frac{12.3}{123} = 0.1m$$

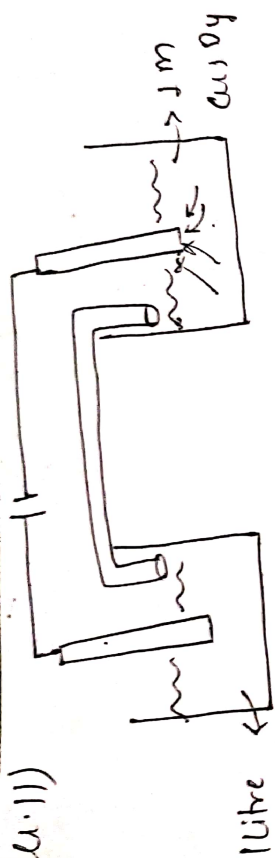
$$1 \text{ mole of N-benzene} = 6F$$

$$0.1 \text{ mole of N-benzene} = 0.6F$$

$$= 0.6 \times 96500$$

$$Q = 57900C$$

Qu. 11)



1M = 1 mole of CuSO_4 in 1 litre of solution

$$\text{mole} = 1 \times 63.5 \text{ g} = w$$

$$w = \frac{At \times I \times t}{\text{val} \times F}$$

$$63.5 = \frac{63.5 \times 1 \times t}{2 \times 96500}$$

$$t = 53.61 \text{ hr}$$

$$w = 50 \text{ g}$$

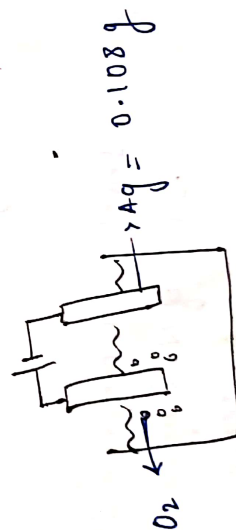
$$w = \frac{At \times I \times t}{\text{val} \times F}$$

$$50 = \frac{63.5 \times 1 \times t}{2 \times 96500}$$

$$t = \frac{50 \times 2 \times 96500}{63.5 \times 3600}$$

$$t = 41.24 \text{ hr}$$

For Zn



Qu. 12

$$E_1 = \frac{108}{1}$$

$$E_2 = \frac{32}{4} = 8 \text{ g}$$

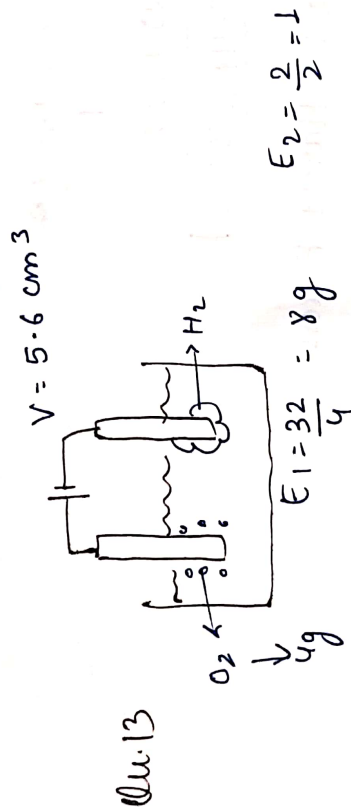
$$\frac{w_1}{w_2} = \frac{E_1}{E_2}$$

$$\frac{0.108}{w_2} = \frac{108}{8}$$

$$w = 8 \times 10^{-3} \text{ g}$$

~~mol~~ mass $\rightarrow$ mole $\rightarrow$ Vol
 $\frac{0}{0} m_w$ $\times 22.4$

$$\frac{18 \times 10^{-3}}{32} = \frac{1}{4} \times 22.4$$



$$\frac{w_1}{w_2} = \frac{E_1}{E_2} = \frac{4}{1} = \frac{8}{1}$$

$$w_2 = \frac{1}{2} = 0.5g$$

mass $\rightarrow$ mole $\rightarrow$ volume
 $\frac{0}{0} m_w$ $\times 22.4$

$$\frac{0.5 \times 22.4 \text{ litre}}{2} = 5.6 \text{ litre}$$