ULTIMATE KCET

CRASH COURSE 2026

Chemistry | Lecture - 01

Electrochemistry

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Recap *of previous lecture*

1. Redox reaction

Topics *to be covered*

1. Electrochemistry

Electrochemistry

$[H^+] = [C\alpha]$

Year	Topic		Number of questions
2025	**Galvanic cell statement type, electronic conductance statement type, Faraday's law - 03**		
2024	Faraday's first law – 02, Debye Onsager equation		03
2023	pH calculations, product of electrolysis – Brine solution , Nernst equation,		03
2022	Gibbs free energy and spontaneity, Molar conductance (02), Catalyst used in Fuel cell, Nernst equation		04
2021	Nernst equation, cell constant, equilibrium constant		03
2020	Debye Onsager equation, Equilibrium constant, Product of electrolysis – NaF		04
2019	Displacement reaction, Emf calculations, Product of electrolysis, Faradays first law of electrolysis		04
2018	Molar conductance/specific conductance, Conductor/Non conductor		03
2017	Product of electrolysis of Aq NaCl		01
2016	Galvanic cell – Oxidation/reduction Vs anode/cathode, secondary cell, Faraday's law of electrolysis		03

Spontaneous redox reaction

→ Galvanic cell → chemical energy
or Voltaic cell → └→ electrical energy

→ Daniel cell ($Cu + Zn$)

Device which converts electrical energy to nonspontaneous chemical energy is called → Electrolytic cell.

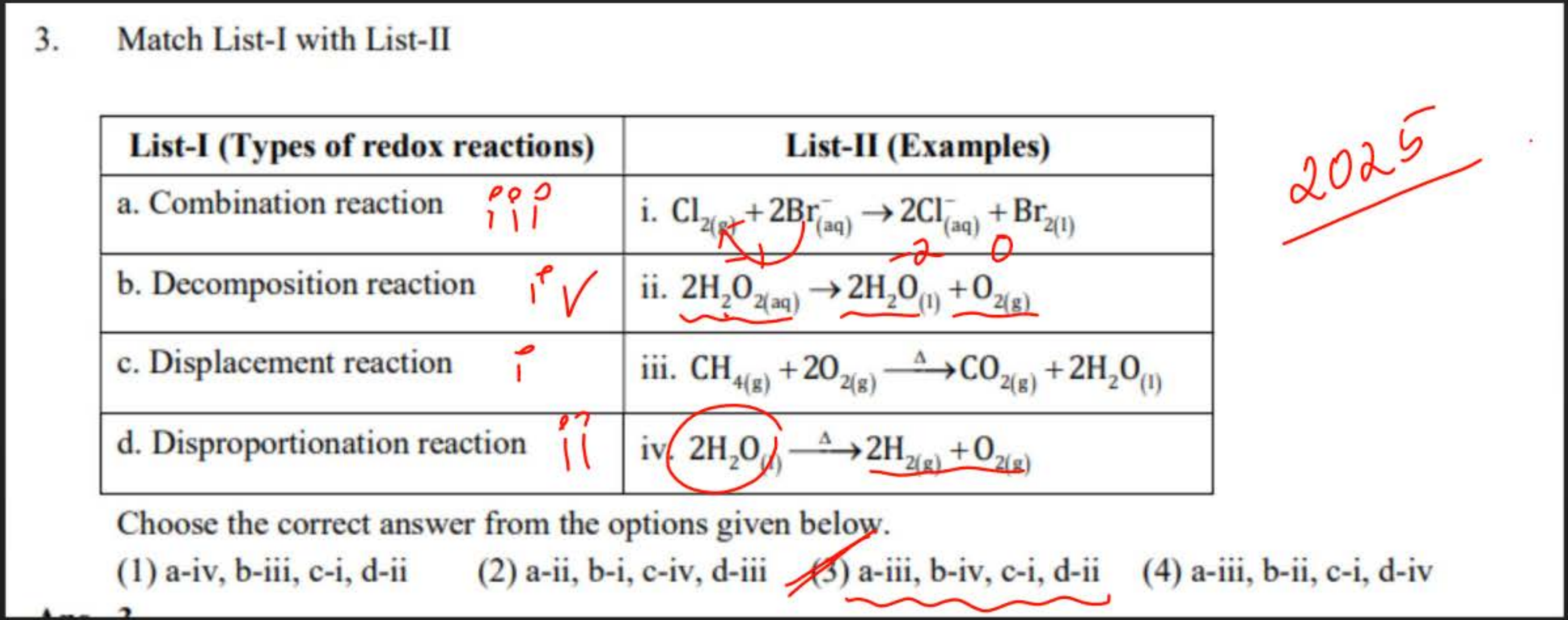
H_2S, SO_2

H_2O_2

$H-O-\overset{-1}{O}-H$

H_2O

$H^+ \ O^{-2}$



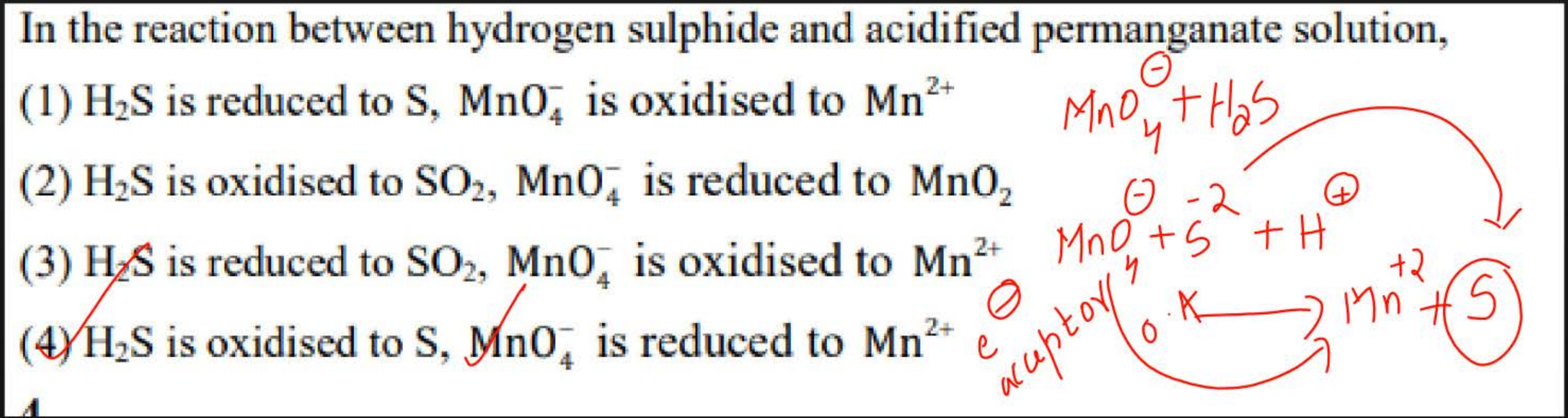
3. Match List-I with List-II

List-I (Types of redox reactions)	List-II (Examples)
a. Combination reaction iii	i. $Cl_{2(g)} + 2Br^-_{(aq)} \rightarrow 2Cl^-_{(aq)} + Br_{2(l)}$
b. Decomposition reaction i ✓	ii. $2H_2O_{2(aq)} \rightarrow \overset{-2}{2H_2O_{(l)}} + \overset{0}{O_{2(g)}}$
c. Displacement reaction i	iii. $CH_{4(g)} + 2O_{2(g)} \xrightarrow{\Delta} CO_{2(g)} + 2H_2O_{(l)}$
d. Disproportionation reaction ii	iv. $2H_2O_{(l)} \xrightarrow{\Delta} 2H_{2(g)} + O_{2(g)}$

Choose the correct answer from the options given below.

(1) a-iv, b-iii, c-i, d-ii (2) a-ii, b-i, c-iv, d-iii ✓(3) a-iii, b-iv, c-i, d-ii (4) a-iii, b-ii, c-i, d-iv

2025



In the reaction between hydrogen sulphide and acidified permanganate solution,

(1) H_2S is reduced to S, MnO_4^- is oxidised to Mn^{2+}

(2) H_2S is oxidised to SO_2, MnO_4^- is reduced to MnO_2

(3) H_2S is reduced to SO_2, MnO_4^- is oxidised to Mn^{2+}

✓(4) H_2S is oxidised to S, MnO_4^- is reduced to Mn^{2+}

$MnO_4^{\ominus} + H_2S$

$MnO_4^{\ominus} + S^{-2} + H^{\oplus} \rightarrow Mn^{+2} + S$

$e^{\ominus}$ acceptor

O.A

O.A

$F^{\ominus}$

$Cl^{\ominus}$

$Br^{\ominus}$

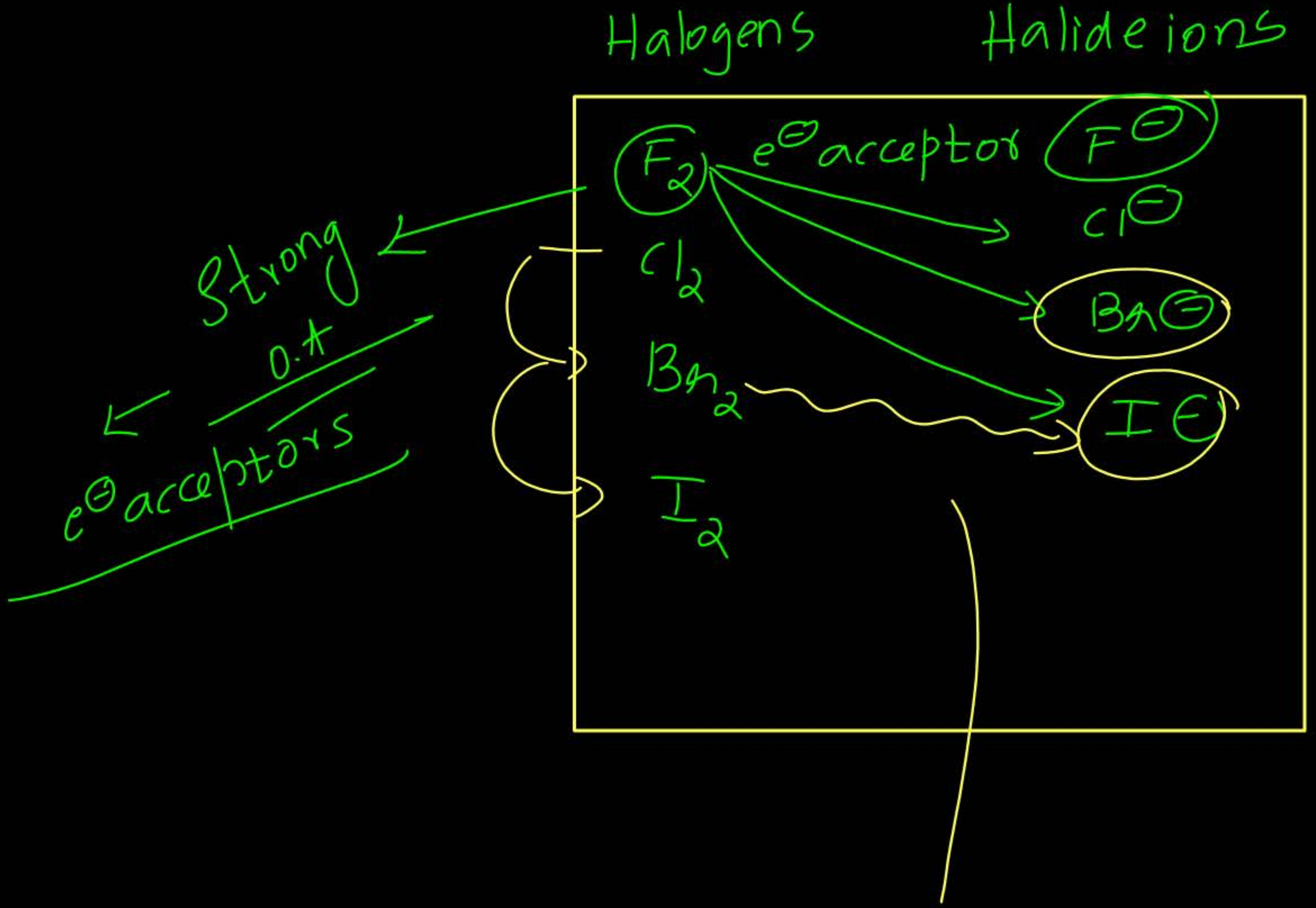
$I^{\ominus}$

O.A

Oxidizing ability

$Br_2 + Cl^{\ominus} \longrightarrow$ No reactions.

$Br_2 + I^{\ominus} \longrightarrow 2Br^{\ominus} + I_2$.



$Cl_2 + F^{\ominus} \longrightarrow$ No reaction

$Cl_2 + 2Br^{\ominus} \longrightarrow 2Cl^{\ominus} + \boxed{Br_2}$

$Br^{\ominus}$

$Br^{\ominus}$

$n_{N_2} = 7.16\times10$

49. The correct statement/s about Galvanic cell is/are

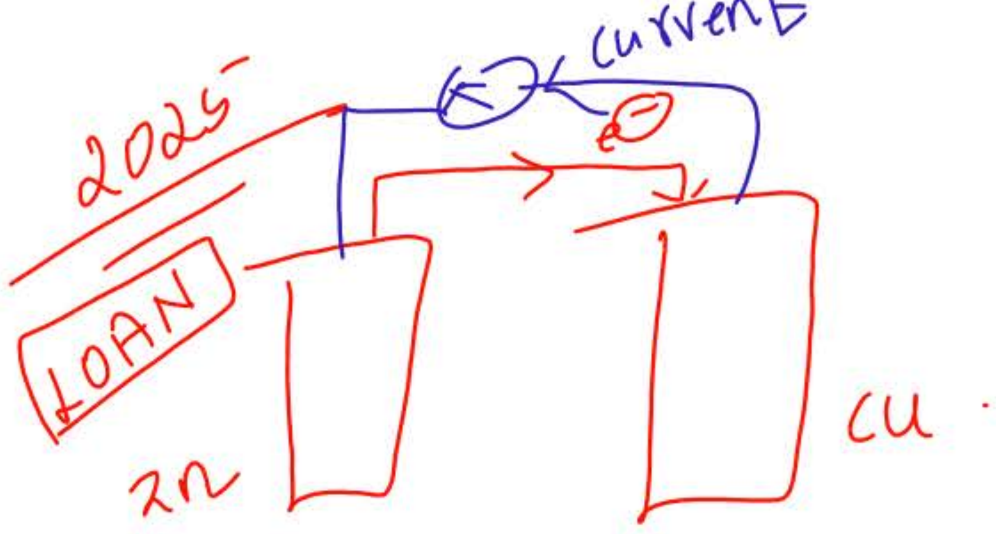
(a) Current flows from cathode to anode

(b) Anode is positive terminal

(c) If $E_{cell} < 0$, then it is spontaneous reaction

(d) Cathode is positive terminal

(1) a and b only (2) a, b and c (3) a and d only (4) b only



50. The electronic conductance depends on

(1) Nature of electrolyte added → ionic conductance (2) The number of valence electrons per atom

(3) Concentration of the electrolyte (4) Size of the ions

Ans. 2

$$\boxed{\begin{array}{l}\Delta G^\circ = -nFE^\circ \\ E^\circ = +ve\end{array}} \rightarrow \text{spontaneous}$$

51. For a given half cell, $\boxed{Al^{3+} + 3e^- \rightarrow Al}$ on increasing of aluminium ion, the electrode potential will

(1) Decrease (2) No change (3) First increase then decrease (4) Increase

Ans. 4

Solution : $Al^{3+} + 3e^- \rightarrow Al(s)$

$$\log\frac{1}{[Al^{3+}]} = -\log[Al^{3+}]$$

✓ $E_{Red} = E^{o}_{Red} - \frac{0.0591}{3}\log\frac{[Al(s)]}{[Al^{3+}]}$

cal cell

$E_{Red} = E^{o}_{Red} - \frac{0.0591}{\underset{3}{\cancel{4}}}\log\frac{\cancel{1}}{[Al^{3+}]}$ [Since active mass of solid = 1]

✓ $E_{Red} = E^{o}_{Red} + \frac{0.0591}{3}\log[Al^{3+}]$ ✓✓

So $E_{Red} \propto [Al^{3+}] \propto$ concⁿ of Al^{3+}

52. Match the following select the correct option for the quantity of electricity, in Cmol^{-1} required to deposit various metals at cathode

96,500 C ≅ 1 Lac

2025

	List – I		List- II
a	Ag^{+} 1F	i	$386000\,\text{Cmol}^{-1}$ → 4F
b	Mg^{2+} 2F	ii	$289500\,\text{Cmol}^{-1}$ → 3F
c	Al^{3+} 3F	iii	$96500\,\text{Cmol}^{-1}$ 1F
d	Ti^{4+} 4F	iv	$193000\,\text{Cmol}^{-1}$ → 2F

(1) a – ii, b – i, c – iv, d – iii

~~(2)~~ a – iii, b – iv, c – ii, d – i

(3) a – iv, b – iii, c – i, d – ii

(4) a – i, b – ii, c – iii, d – iv

Daniell cell

LOAN

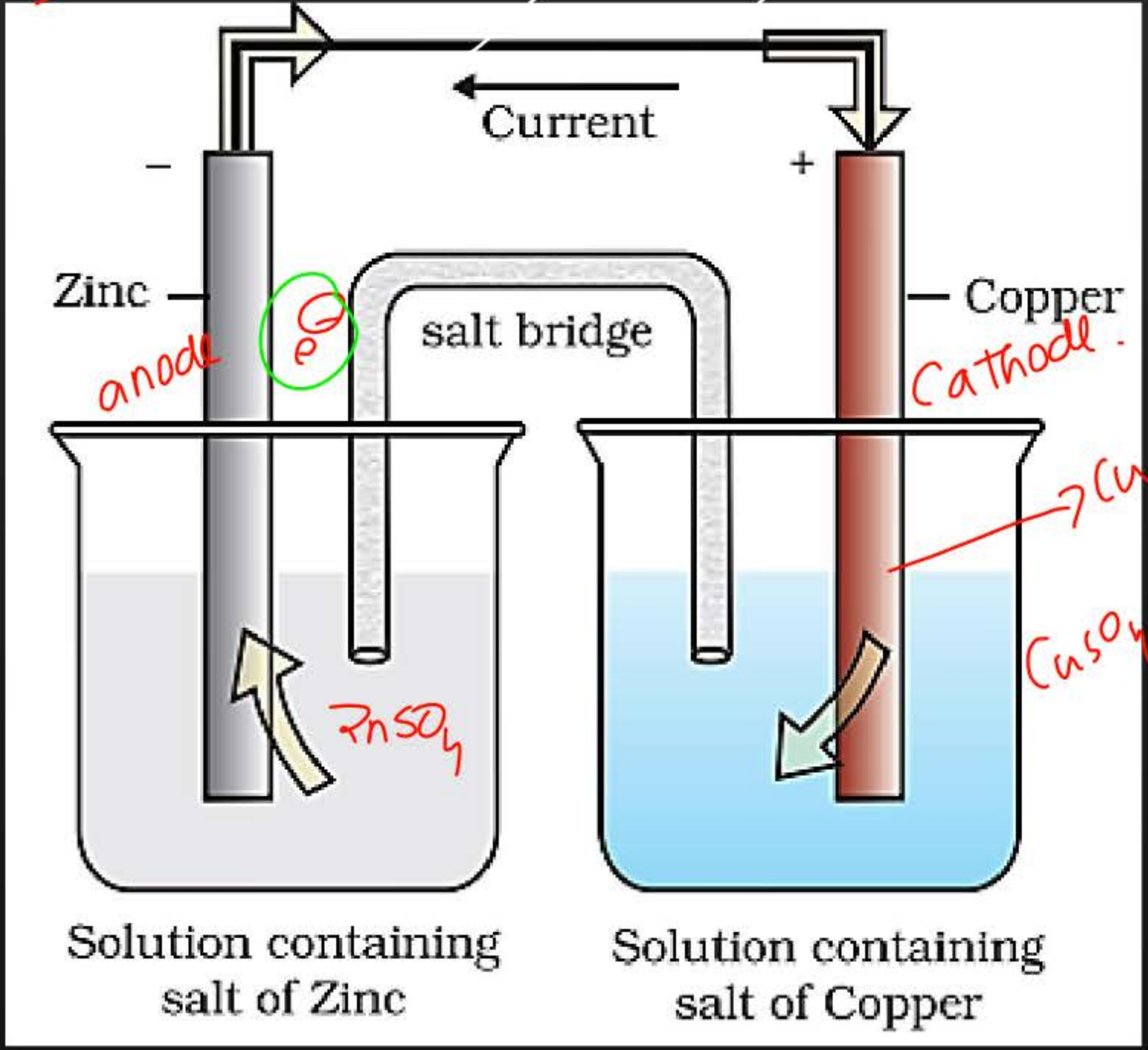
Left oxidation

Anode, -ve

RRCP

[O] → $Zn \longrightarrow Zn^{2+} + 2e^{\ominus}$

[R] $Cu^{2+} + 2e^{\ominus} \longrightarrow Cu\downarrow$



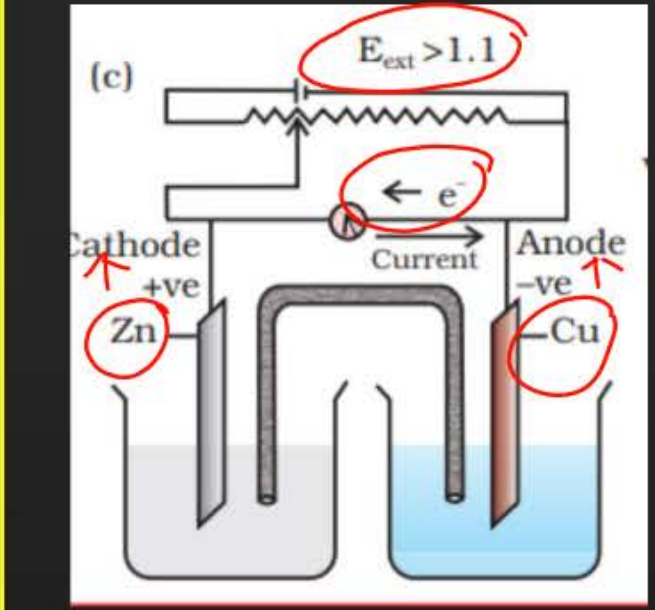
- This cell converts the chemical energy liberated during the redox reaction to electrical energy and has an electrical potential equal to 1.1 V.

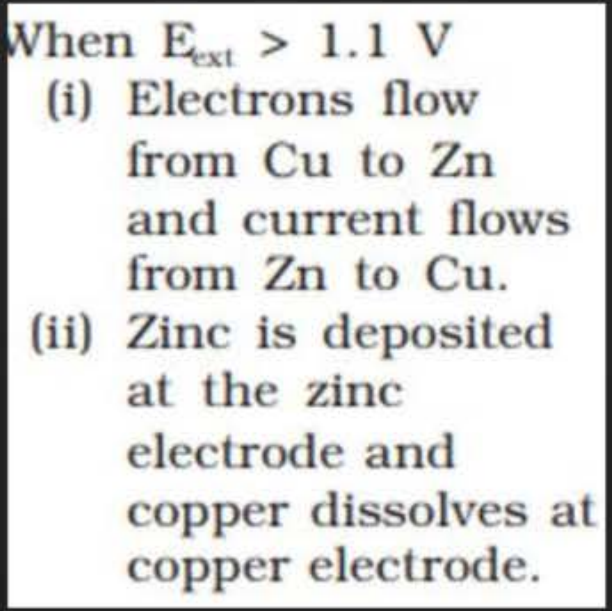
- when concentration of Zn^{2+} and Cu^{2+} ions is unity. Such a device is called a galvanic or a voltaic cell

$$EMF = E^{\circ}_{cathode} - E^{\circ}_{anode}$$

→ +ve value

$e^{\ominus}$
anode → cathode
Current
cathode → anode

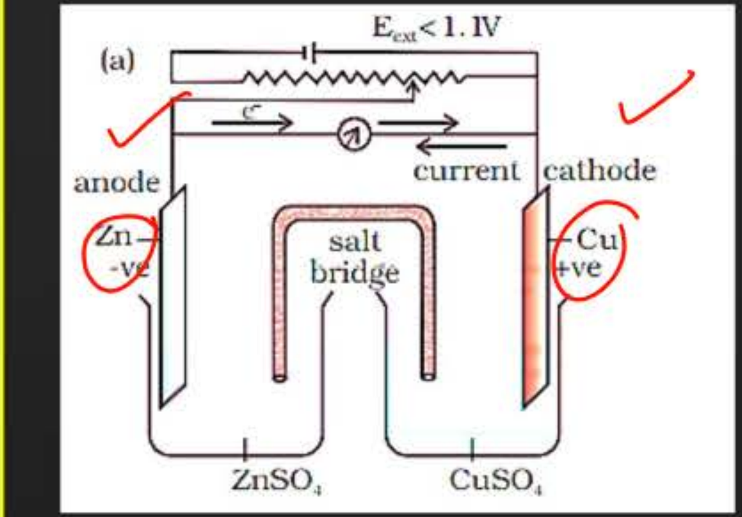


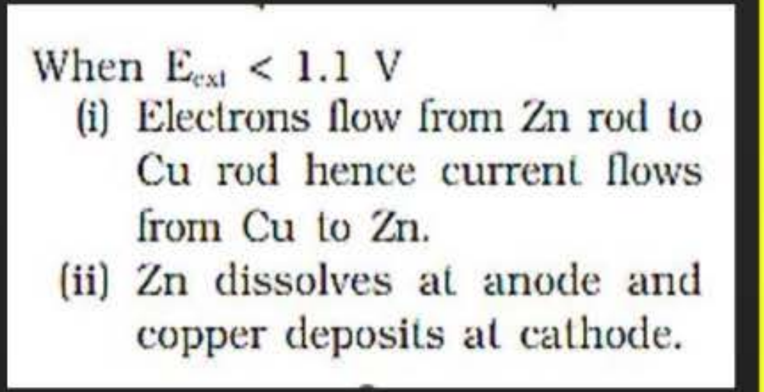

When $E_{ext} > 1.1$ V

(i) Electrons flow from Cu to Zn and current flows from Zn to Cu.

(ii) Zinc is deposited at the zinc electrode and copper dissolves at copper electrode.

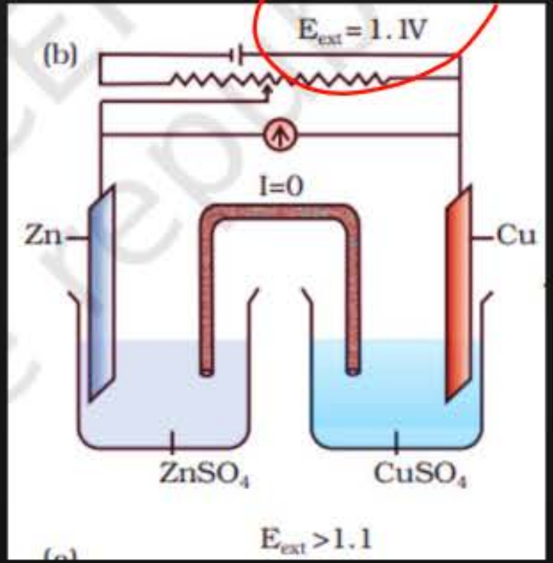
$E^0 = 1.1V$

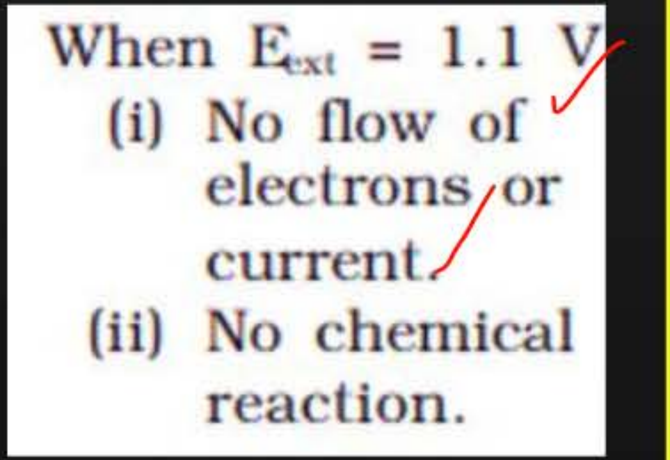



When $E_{ext} < 1.1$ V

(i) Electrons flow from Zn rod to Cu rod hence current flows from Cu to Zn.

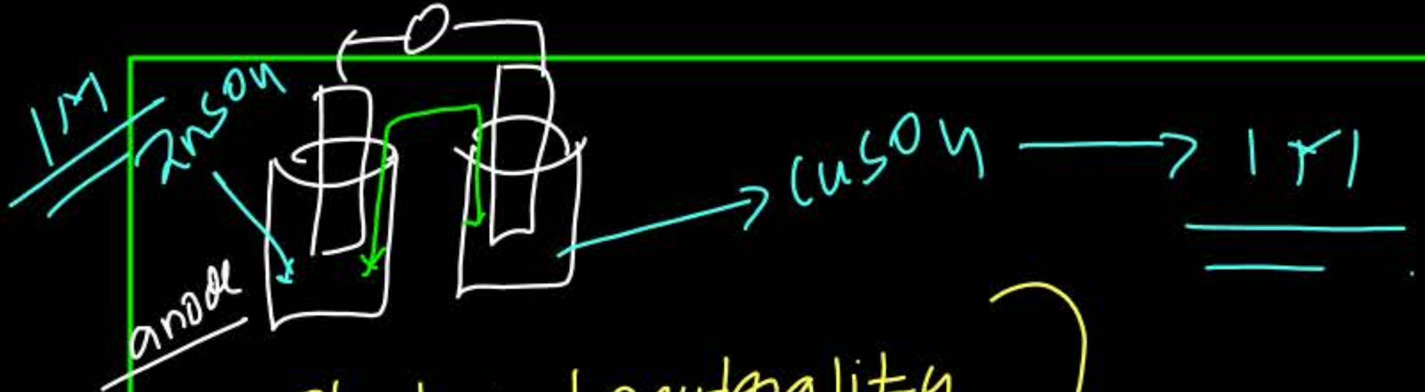
(ii) Zn dissolves at anode and copper deposits at cathode.




When $E_{ext} = 1.1$ V

(i) No flow of electrons or current.

(ii) No chemical reaction.



→ Electrical neutrality

→ inner circuit connection. } salt bridge function

→ Cell potential → / EMF →

$$\boxed{\begin{aligned} E^\circ &= E^\circ_{RHS} - E^\circ_{LHS} \\ E^\circ &= E^\circ_{cathode} - E^\circ_{anode} \end{aligned}}$$

→ $E_{cell} = E^\circ_{cell} - \frac{0.0591}{n}\log\frac{[\text{anode ion}]}{[\text{cathode ion}]}$

$$\overset{\curvearrowright}{Zn \mid Zn^{2+}} \parallel Ag^{+} \mid Ag \longrightarrow \text{cathode.}$$

anode

anode electrolyte

Cathode electrolyte

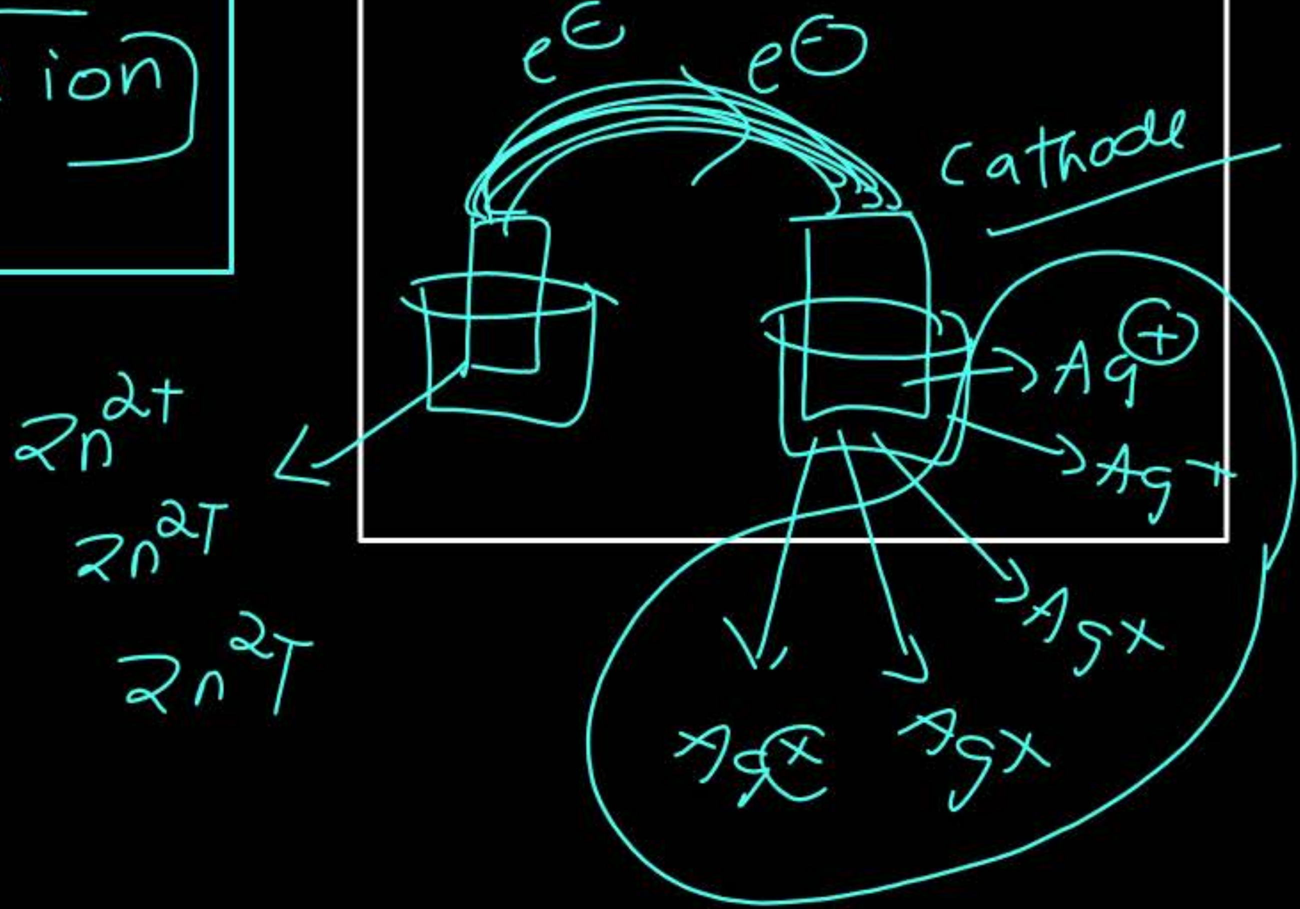
$$E_{cell} = E^{\circ}_{cell} - \frac{0.059}{n} \log \frac{[\text{anode ion}]}{[\text{cathode ion}]}$$

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0591}{2} \log \frac{[Zn^{2+}]}{[Ag^{+}]}$$

E_{cell} = is maximum

when cathode ion

conc is max.



Reduction reaction

O.A.

e⊖ acceptor

oxidant

$$Zn^{2+} + 2e^{\ominus} \longrightarrow Zn_{(s)}$$

→ Reductant

→ R.A

→ e⊖ rich.

$$\boxed{E_{RP} = E^{\circ}_{cell} - \frac{0.0591}{2}\log\frac{[Zn]_{(s)}}{[Zn^{2+}]}}$$

$$\rightarrow \boxed{E^{\circ}_{RP} = E^{\circ}_{cell} - \frac{0.0591}{n}\log\frac{[\text{Reductant}]}{[\text{oxidant}]}.}$$

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

Oxidation

$$Zn \longrightarrow Zn^{2+} + 2e^{\ominus}$$

Zn ↓ Reductant ↓ e⊖ donar ↓ R.A.

Zn^{2+} ↓ oxidant ↓ e⊖ acceptor ↓ O.A

→ ↓ Oxidation potential

$$E_{op} = E^{\circ}_{op} - \frac{0.0591}{n}\log\frac{Zn^{2+}}{Zn}.$$

$$= E^{\circ}_{op} - \frac{0.0591}{n}\log\frac{\text{oxidant}}{\text{Reductant}}.$$

→ E_{cell} for diff. conc.

$E^0_{cell} = E^0_{cathode} - E^0_{anode}$.

$$\boxed{\Delta G^0 = -2.303\, nRT \log K_c.}$$

$E_{cell} = +ve$

spontaneous → anode to cathode → $e^\ominus$ movement.

$$\boxed{E_{cell} = \text{zero at equilibrium}}$$

Gibb's free energy

$$\therefore \boxed{E^0_{cell} = \frac{0.0591}{n} \log K_c}$$

$$\boxed{\Delta G^0 = -nFE^0}$$

n → no. of moles of $e^\ominus$ involved in redox reaction

F → 96,500 C, $E^0 = E^0_{RHS} - E^0_{LHS}$

$E^0_{cathode} - E^0_{anode}$.

$aA + bB \rightleftharpoons cC + dD$

$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$

$\Delta G^0 = -ve$ → for any redox reactions to be spontaneous.

→ Electrolytic cell

$$\boxed{R \propto \frac{1}{a}} \longrightarrow R = \rho \frac{l}{a}$$

→ ohm

Ω

ρ → resistivity

specific resistance.

$$\boxed{\rho = R \times \frac{a}{l}}$$

$$\rho = ohm\, \frac{\cancel{m^2}}{\cancel{m}}$$

$$\boxed{\rho = ohm\, m}$$

$$\boxed{R = \frac{1}{G}}$$

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

G = conductance

R = Resistance.

Kappa → specific conductance.

[conductivity]

$$\boxed{K = \frac{1}{\rho}}$$

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

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

$$K = \frac{1}{R} \times G^*$$

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

unit for K

$K = ohm^{-1} m^{-1}$

$ohm^{-1} cm^{-1}$

$S\, cm^{-1}$ ✓

$S\, m^{-1}$

Molar conductance

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

K → specific conductance

C → concentration

$$\Lambda_m = \frac{1000 \times S\, \cancel{cm^{-1}}}{mol\, \cancel{cm^{-3}}} \quad cm^{-2}$$

$$\Lambda_m = S\, cm^2 mol^{-1}$$

$C = mol\, dm^{-3}$

$C = mol\, L^{-1}$

$C = mol\, cm^{-3}$

Kohlrausch's law

$$\boxed{\Lambda_m^{\infty} = x\lambda_m^{\infty +} + y\lambda_m^{\circ\ominus}}$$

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

$$\boxed{[H^+] = c\alpha}$$

↓

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

or

$pH = \log\frac{1}{[H^+]}$

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

Faraday's first law

$W \propto Q$.

$Q = It$

↳ As.

$W \propto It$

$$\boxed{W = zIt}$$

z → electrochemical equivalent

→ amount of substance deposited when 1 ampere current is passed for 1 second.

$$\boxed{z = \frac{\text{Equivalent weight}}{96{,}500}}$$

$^{27}_{13}Al$

Al^{+3}

$$\text{Equivalent weight} = \frac{\text{Atomic mass}}{\text{charge}} = \frac{\overset{9}{\cancel{27}}}{\cancel{3}} = 9 \ g_{eq}.$$

Equivalent weight

of Ag $= \frac{108}{1} = 108$ g eq wt.

^{108}Ag

Ag^{+1}

$CuSO_4 \longrightarrow Cu^{2+} \longrightarrow$ copper.

$Cu_2O \longrightarrow Cu^{+1} \longrightarrow$ cuprous

$$\boxed{\text{eq.wt of } Cu^{2+} = \frac{\overset{31.}{63.5}}{2} = \underline{\underline{31.7\,g}}}$$

$^{63.5}Cu$

Eq.wt of oxygen

$^{16}_{8}O, \; O^{-2}$

$Eq.wt = \frac{16}{8} = 8\,g$ eq.wt

faraday's second law

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

Same Q $\rightarrow CuSO_4$

$\rightarrow ZnSO_4$

Stronger oxidizing agent

Weaker reducing agent

$F_2(g) + 2e^-$	$\longrightarrow 2F(aq)$	2.87
$H_2O_2(aq) + 2H^+(aq) + 2e^-$	$\longrightarrow 2H_2O(l)$	1.78
$MnO_4^-(aq) + 8H^+(aq) + 5e^-$	$\longrightarrow Mn^{2+}(aq) + 4H_2O(l)$	1.51
$Cl_2(g) + 2e^-$	$\longrightarrow 2Cl^-(aq)$	1.36
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^-$	$\longrightarrow 2Cr^{3+}(aq) + 7H_2O(l)$	1.33
$O_2(g) + 4H^+(aq) + 4e^-$	$\longrightarrow 2H_2O(l)$	1.23
$Br_2(l) + 2e^-$	$\longrightarrow 2Br^-(aq)$	1.09
$Ag^+(aq) + e^-$	$\longrightarrow Ag(s)$	0.80
$Fe^{3+}(aq) + e^-$	$\longrightarrow Fe^{2+}(aq)$	0.77
$O_2(g) + 2H^+(aq) + 2e^-$	$\longrightarrow H_2O_2(aq)$	0.70
$I_2(s) + 2e^-$	$\longrightarrow 2I^-(aq)$	0.54
$O_2(g) + 2H_2O(l) + 4e^-$	$\longrightarrow 4OH^-(aq)$	0.40
$Cu^{2+}(aq) + 2e^-$	$\longrightarrow Cu(s)$	0.34
$Sn^{4+}(aq) + 2e^-$	$\longrightarrow Sn^{2+}(aq)$	0.15
$2H^+(aq) + 2e^-$	$\longrightarrow H_2(g)$	0
$Pb^{2+}(aq) + 2e^-$	$\longrightarrow Pb(s)$	−0.13
$Ni^{2+}(aq) + 2e^-$	$\longrightarrow Ni(s)$	−0.26
$Cd^{2+}(aq) + 2e^-$	$\longrightarrow Cd(s)$	−0.40
$Fe^{2+}(aq) + 2e^-$	$\longrightarrow Fe(s)$	−0.45
$Zn^{2+}(aq) + 2e^-$	$\longrightarrow Zn(s)$	−0.76
$2H_2O(l) + 2e^-$	$\longrightarrow H_2(g) + 2OH^-(aq)$	−0.83
$Al^{3+}(aq) + 3e^-$	$\longrightarrow Al(s)$	−1.66
$Mg^{2+}(aq) + 2e^-$	$\longrightarrow Mg(s)$	−2.37
$Na^+(aq) + e^-$	$\longrightarrow Na(s)$	−2.71
$Li^+(aq) + e^-$	$\longrightarrow Li(s)$	−3.04

Weaker oxidizing agent

Stronger reducing agent

cathode

anode ←

$e^{\ominus}$ acceptor.

O.A

→ cathode

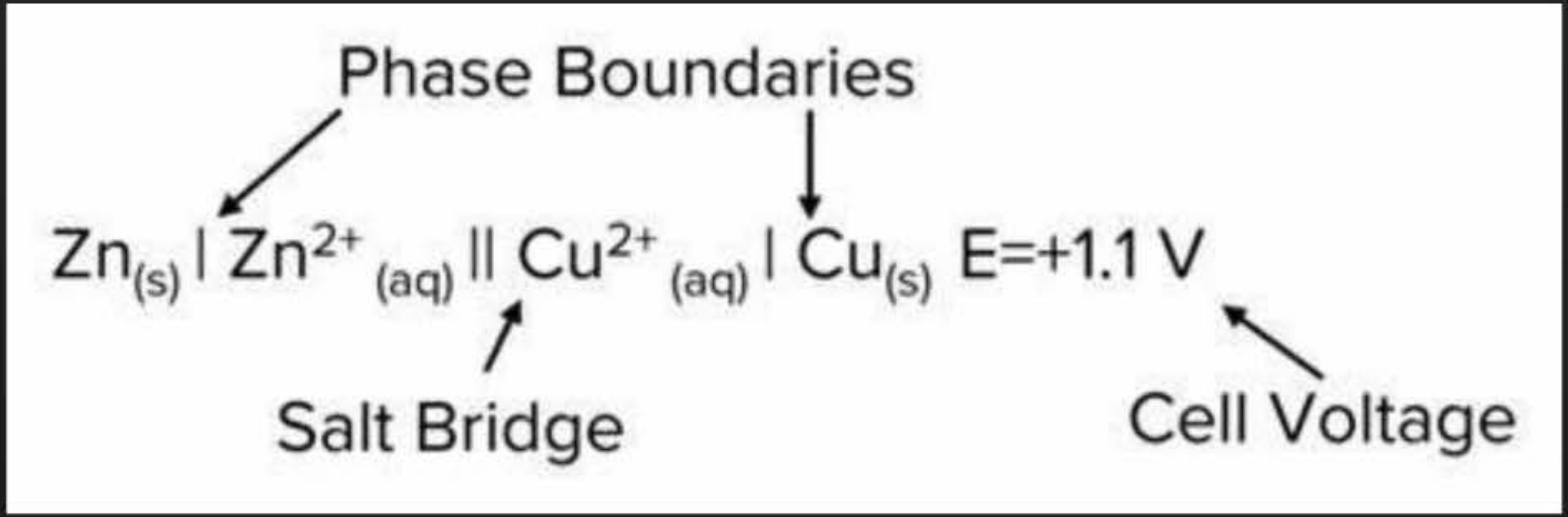
R.A

anode

$e^{\ominus}$ donors

Representation of Electrochemical Cell

An electrochemical cell is represented in a manner as illustrated below for the Daniell cell.



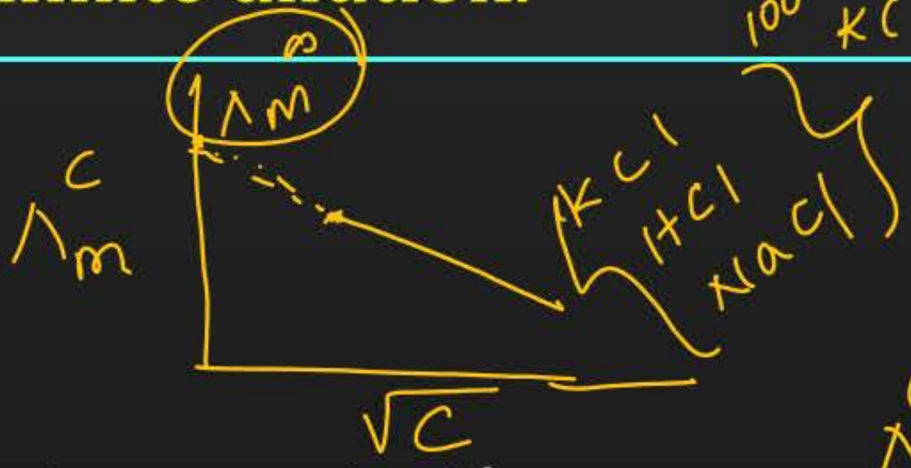
$$Zn_{(s)} \mid Zn^{2+}{}_{(aq)} \parallel Cu^{2+}{}_{(aq)} \mid Cu_{(s)} \quad E=+1.1\ V$$

- By convention, the electrode on which oxidation takes place is written on the left hand side.
- The other electrode on which reduction takes place is written on the right hand side.

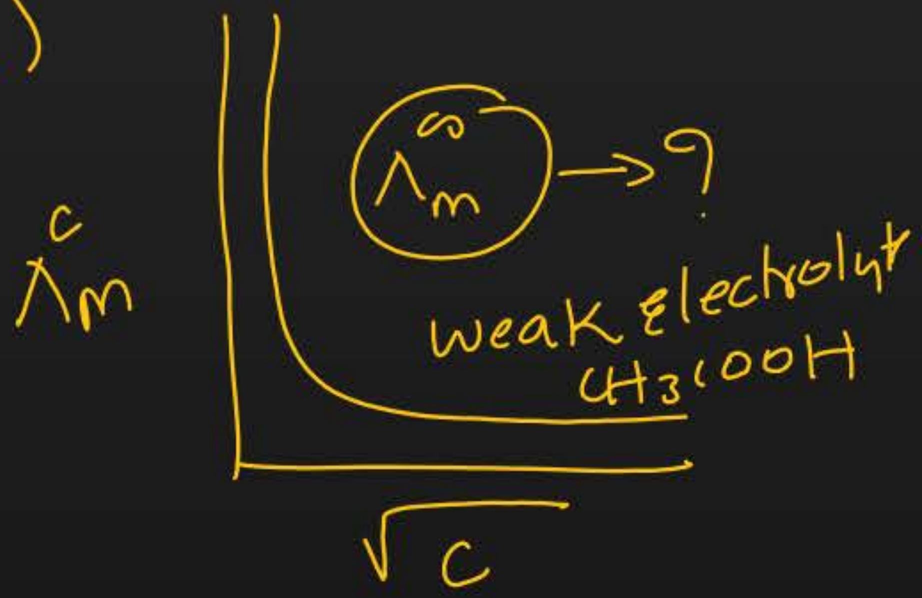
Molar conductance ↑ with dilution.

Debye Huckel has given the following equation to co-relate the molar conductance at any concentration and molar conductance at infinite dilution.

100 KCl ⟶ 100 K^+ + 100 Cl^-



$$\boxed{\Lambda_m^c = \Lambda_m^\infty - A\sqrt{C}}$$



Here, Λ_m^c = Molar conductivity at concentration c.

Λ_m^∞ or Λ_m^c = Molar conductivity at infinite dilution

C = Molar concentration of the electrolyte.

A = Constant depending upon the nature of solvent and temperature.

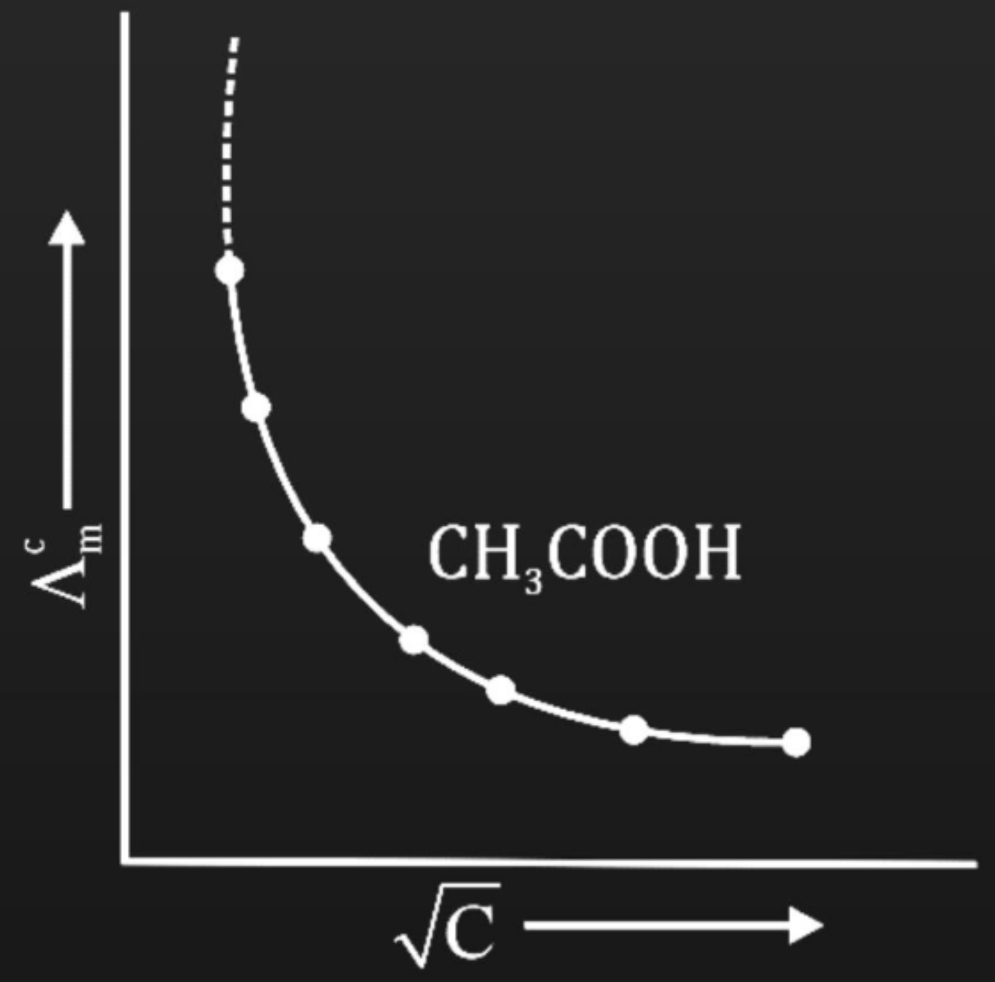


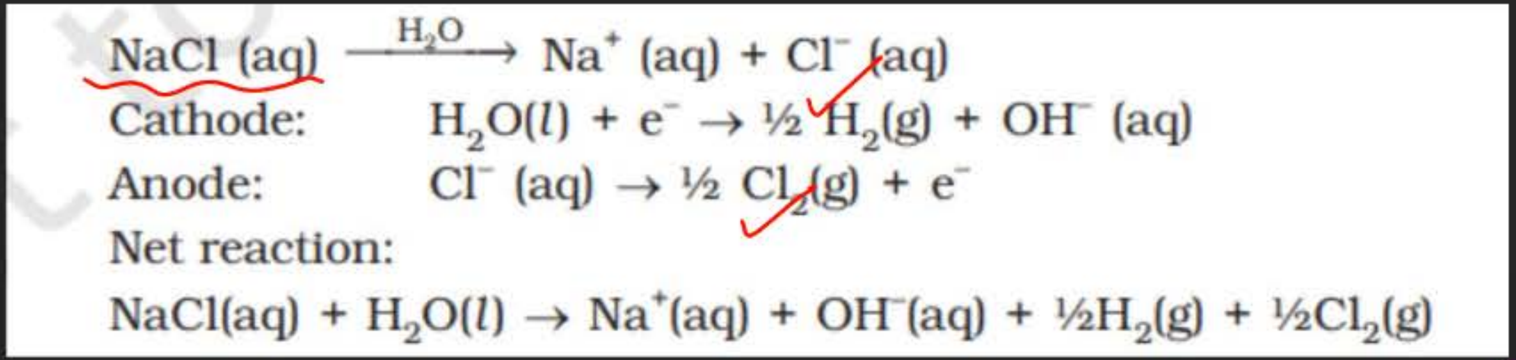
Figure 3.17 variation of Λ_m^c with $\sqrt{C}$ for weak electrolyte (CH_3COOH)

$K^+ Cl^{-1} \rightarrow 1:1$ $Ca^{+2} Cl^{-1}$ $Na^{\oplus} SO_4^{2-}$

- For a particular solvent at a given temperature, the value of A depends upon nature of the electrolyte i.e. charges on the cation and anion produced on the dissociation of the electrolyte.
- For example, KCl, $CaCl_2$ and Na_2SO_4 are known as 1-1, 2-1 and 1-2 electrolytes respectively.

+1 −1 +1 −1 +1 −1

- All the electrolytes of the same type (e.g. NaCl, KCl, KNO_3) have the same value of A. These are 1-1 electrolytes.
- A graph between Λ_m and $\sqrt{C}$ for a strong electrolyte (KCl) has been shown in the Fig. 3.16.
- By extrapolation, the limiting molar conductance (Λ_m^0) can be obtained.



$NaCl\,(aq) \xrightarrow{H_2O} Na^+\,(aq) + Cl^-\,(aq)$

Cathode: $H_2O(l) + e^- \rightarrow \frac{1}{2} H_2(g) + OH^-\,(aq)$

Anode: $Cl^-\,(aq) \rightarrow \frac{1}{2} Cl_2(g) + e^-$

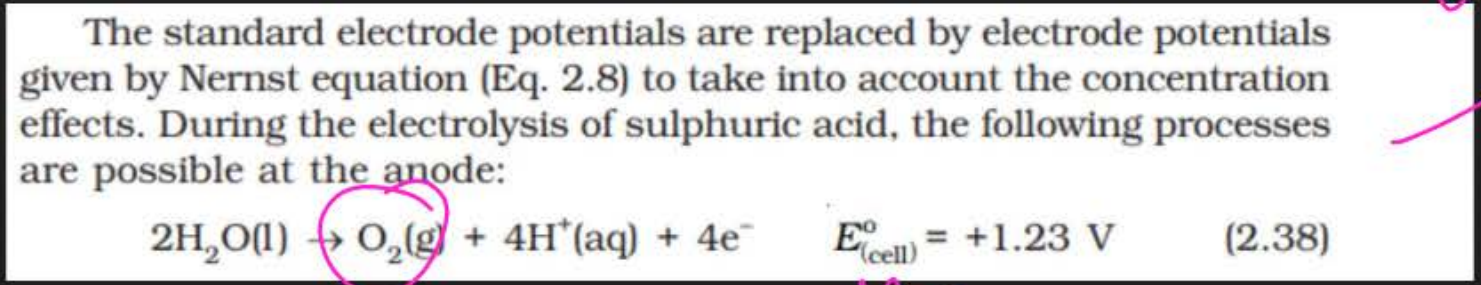
Net reaction:

$NaCl(aq) + H_2O(l) \rightarrow Na^+(aq) + OH^-(aq) + \frac{1}{2}H_2(g) + \frac{1}{2}Cl_2(g)$

anode → O_2

cathode → H_2

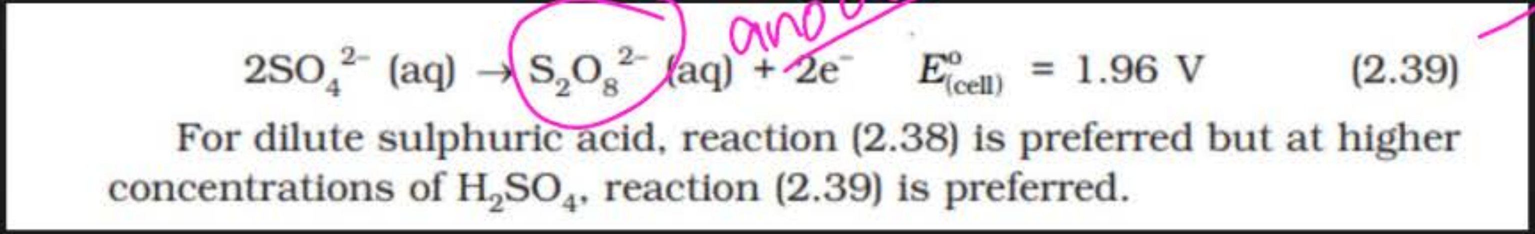
→ dil. H_2SO_4



The standard electrode potentials are replaced by electrode potentials given by Nernst equation (Eq. 2.8) to take into account the concentration effects. During the electrolysis of sulphuric acid, the following processes are possible at the anode:

$$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^- \qquad E^o_{(cell)} = +1.23\ V \qquad (2.38)$$

→ conc. H_2SO_4



$$2SO_4^{2-}\,(aq) \rightarrow S_2O_8^{2-}\,(aq) + 2e^- \qquad E^o_{(cell)} = 1.96\ V \qquad (2.39)$$

anode

For dilute sulphuric acid, reaction (2.38) is preferred but at higher concentrations of H_2SO_4, reaction (2.39) is preferred.

cathode → $H^{\oplus}$ → $H_2\uparrow$.

Question

$Zn \longrightarrow Zn^{2+} + 2e^-$: $E^\circ = +\overset{\text{cathode}}{0.76}$ V, $Fe \longrightarrow Fe^{2+} + 2e^-$; $E^\circ = +\overset{\text{anode}}{0.41}$ V The e.m.f. for the cell reaction $Fe^{2+} + Zn \longrightarrow Zn^{2+} + Fe$ is

- (A) -0.35 V ✗
- [✓ (B)] +0.35 V
- (C) +1.17 V
- (D) -1.17 V ✗

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}.$$

$$= 0.76 - 0.41$$

$$\boxed{E^\circ_{cell} = 0.35\,V}$$

Question

For a cell the reaction is $Mg_{(s)} + Cu^{2+}_{(aq)} \rightarrow Cu_{(s)} + Mg^{2+}_{(aq)}$ If the standard reduction potentials of Mg and Cu are -2.37 V and 0.34 V respectively, the EMF of the cell is

A) –2.71 V

[✓ B)] 2.71 V

C) –2.03 V

D) 2.03 V

-2.37 → anode

0.34 → cathode

$$\begin{array}{r} \overset{1}{2}.37 \\ 0.34 \\ \hline 2.71 \end{array}$$

$$E_{cell} = \boxed{E^\circ_{cathode} - E^\circ_{anode}}$$

$$= 0.34 - (-2.37)$$

$$= 0.34 + 2.37$$

$e^\ominus$ donar
↓
R.A

if E° value is -ve or less positive ⟶ anode

if E° value is +ve or more +ve ⟶ cathode

($e^\ominus$ acceptor)
↓
O.A.

Question

Which is the correct representation for Nernst equation ?

A $E_{RP} = E^{0}_{RP} - \frac{0.059}{n}\log\frac{[\text{oxidant}]}{[\text{reductant}]}$

B ✓ $E_{OP} = E^{\circ}_{OP} - \frac{0.059}{n}\log\frac{[\text{oxidant}]}{[\text{reductant}]}$

C $E_{OP} = E^{\circ}_{OP} - \frac{0.059}{n}\log\frac{[\text{reductant}]}{[\text{oxidant}]}$

D $E_{RP} = E^{\circ}_{RP}$ ✗ $- \frac{0.59}{n}\log\frac{[\text{reductant}]}{[\text{oxidant}]}$

$Zn^{2+} + 2e^{\ominus} \longrightarrow Zn \longrightarrow$ Reduction

$E_{RP} = E^{\circ}_{RP} - \frac{0.059}{n}\log\frac{[Zn]}{[Zn^{2+}]}$ $[Zn]$ → reductant, $[Zn^{2+}]$ → oxidant.

$Zn \longrightarrow Zn^{2+} + 2e^{\ominus} \longrightarrow$ oxidation

$E_{op} = E^{\circ}_{op} - \frac{0.0591}{n}\log\frac{[Zn^{2+}]}{[Zn]}$ $[Zn^{2+}]$ → oxidant, $[Zn]$ → Reductant

Question

For the reaction $2A_{(s)} + B^{2+}_{(aq)} \longrightarrow 2A^{+}{}_{(aq)} + \cancel{B}_{(s)}$, Nernst equation for the EMF of the cell is

298K

A $E = E^{o} - \frac{RT}{2F}\ln\frac{[A^{+}]^{2}}{[B^{2+}]}$ ✓

B $E = E^{\circ} + \frac{RT}{2F}\ln\frac{[A^{+}]^{2}}{[B^{2+}]}$

C $E = E^{o} - \frac{RT}{F}\ln\frac{[A^{+}]^{2}}{\sqrt{[B^{2+}]}}$ ✗

D $E = E^{\circ} + \frac{RT}{F}\ln\frac{[A^{+}]^{2}}{[B^{2+}]}$ ✓

R = 8.314
T = 298K
F = 96,500C

$\frac{0.059}{}$

$$\boxed{E = E^{0}_{cell} - \frac{RT}{2F}\log\frac{[A^{+}]^{2}}{[B^{2+}]}}$$

$$E_{cell} = E^{0}_{cell} + \frac{RT}{2F}\log\frac{[B^{2+}]}{[A^{+}]^{2}}$$

Question

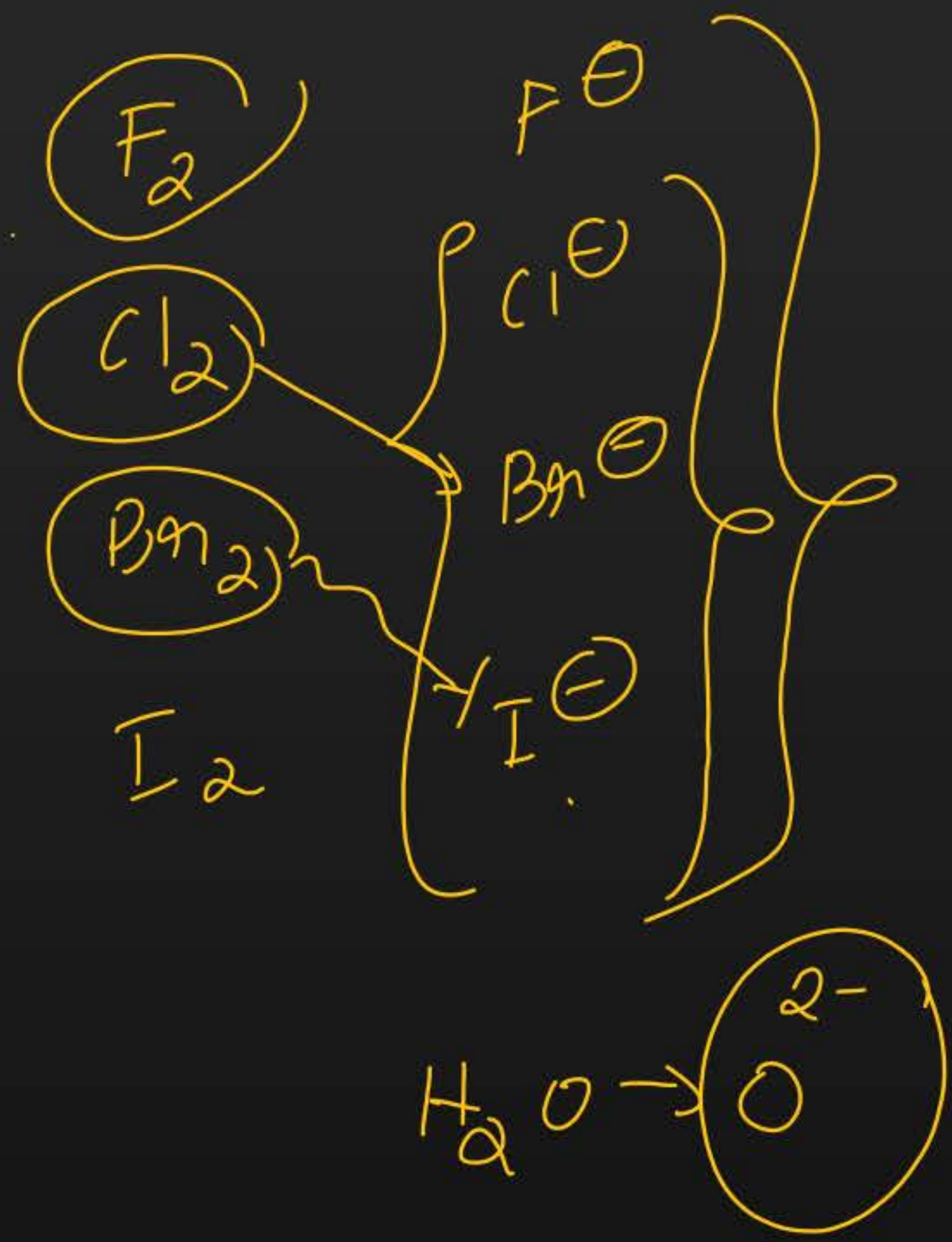
Which reaction is not feasible?

A $2K\overset{\ominus}{I} + Br_2 \rightarrow 2K\overset{\ominus}{Br} + I_2$ ✓

B $2K\overset{\ominus}{Br} + I_2 \rightarrow 2K\overset{\ominus}{I} + Br_2$ ✗ not feasible.

C $2K\overset{\ominus}{Br} + Cl_2 \rightarrow 2K\overset{\ominus}{Cl} + Br_2$ ✓

D $2H_2O + 2\,F_2 \rightarrow 4H\overset{\ominus}{F} + O_2$



Question

The value of $\Delta G(\mathrm{kJ\ mol^{-1}})$ for the cell having $E^\circ_{Cu} = 0.059$ is (take $1F = 96500\ \mathrm{C\ mol^{-1}}$)

- **A** −5.7
- **B** 5.7
- **C** 11.4
- [✓ **D**] −11.4

$\Delta G^\circ = -nFE^\circ$

$\Delta G^\circ = -2\times 0.\underset{\sim}{059}\times 96,500$

$\Delta G^\circ = -\left(2\times 0.\underset{\sim}{06}\times 96,500\right)$

$= -2\times \frac{6}{\cancel{100}}\times 96,5\cancel{00}$

$\Delta G^\circ = \frac{-1,1400\ J}{1000} = -11.4$

$\boxed{Cu^{2+}} + 2e^{\ominus} \longrightarrow Cu$

→ More hydration enthalpy

$n = 2$

$E^\circ = 0.059$

Question

The equilibrium constant (K) for the reaction $Cu_{(s)} + 2Ag^{\circ}_{(4)} \longrightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$ will be [Given $E^{\circ}_{cell} = 0.46$]

- **A** $K_C = \text{Antilog } 15.6$
- **B** $K_C = \text{Antilog } 2.5$
- **C** $K_C = \text{Antilog } 1.5$
- **D** $K_C = \text{Antilog } 12.2$

$E^{\circ}_{cell} = 0.46V$

$n = 2$

$$E^{\circ}_{cell} = \frac{0.0591}{n} \log K_c$$

$$0.46 = \frac{0.059}{2} \log K_C$$

$$\frac{0.46 \times 2}{0.059} = \log K_C$$

$$\frac{0.92}{0.06} = \frac{\overset{154}{\cancel{92}} \times \cancel{100}}{\cancel{100} \times 6}$$

$15.3 = \log K_c \leftarrow$

$K_c = \text{Antilog } (15.3)$

15.3

$\cancel{92}$ .20

6 1

Question

Specific conductance of 0.1 M nitric acid is $6.3 \times 10^{-2} \text{ohm}^{-1} \text{ cm}^{-1}$. The molar conductance of the solution is

- **A** $630 \text{ohm}^{-1} \text{ cm}^2 \text{ mol}^{-1}$
- **B** $315 \text{ohm}^{-1} \text{ cm}^2 \text{ mol}^{-1}$
- **C** $6.300 \text{ohm}^{-1} \text{ cm}^2 \text{ mol}^{-1}$
- **D** $63.0 \text{ohm}^{-1} \text{ cm}^2 \text{ mol}^{-1}$

$$\boxed{\Lambda_m = \frac{1000 \times K}{c}}$$

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

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

$$= 6.3 \times \cancel{10^{-2} \times 10^{4}}$$

$$= 6.3 \times 10^{2} = \boxed{\begin{array}{l} 630 \text{ S cm}^2 \text{ mol}^{-1} \\ 630 \text{ ohm}^{-1} \text{ cm}^2 \text{ mol}^{-1} \end{array}}$$

$$\frac{63 \times \cancel{10^{2}}}{\cancel{10}}$$

Question

$\boxed{Ba^{2+}}$ $\boxed{SO_4^{2-}}$

Molar ionic conductivities of a bivalent electrolyte are 57 and 73. The molar conductivity of the solution will be

A ✓ 130 S cm^2 mol^{-1}

B 65 S cm^2 mol^{-1}

C 260 S cm^2 mol^{-1}

D 187 S cm^2 mol^{-1}

$$\Lambda_m^\infty = x\lambda^0_{2+} + y\lambda^0_{2-}$$

$$\Lambda^0_m = 57 + 73 = 130 \text{ S cm}^2\text{mol}^{-1}$$

Question

How many coulombs are required for the reduction of 1 mol of Cu^{2+} to Cu ?

- A $1.23 \times 10^5 C$ ✗
- B $1.63 \times 10^5 C$
- [✓ C] $1.93 \times 10^5 C$
- D $2.12 \times 10^5 C$ ✗

$Cu^{2+} + 2e^{\ominus} \longrightarrow Cu$

$2F$

Question

At STP 1.12 litre of H_2 is obtained on flowing a current for 965 seconds in a solution. The value of current is

- (A) 10 ✓
- (B) 1.0
- (C) 1.5
- (D) 2.0

Molar volume → 1 mole = 22.4 L

22.4 L → 2 g (H_2)

1.12 L → ?

$$\frac{\cancel{1.12 L} \times \cancel{2g}}{\cancel{22.4 L}} = \boxed{0.1 g}$$

$$\boxed{W = Z \times I \times t}$$

$$H^{\oplus} + e^{\ominus} \longrightarrow H$$

✓ $H = \frac{1}{1} = 1$ g eq.wt

$$0.1 = \frac{\text{Eq wt of H}}{96{,}500} \times I \times 965$$

$$0.1 = \frac{1}{\cancel{96{,}500}} \times I \times \cancel{965}$$

$$0.1 = \frac{I}{100}$$

$$\boxed{I = 10 A}$$

Question

At STP 1.12 litre of H_2 is obtained on flowing a current for 965 seconds in a solution. The value of current is

- **A** 10
- **B** 1.0
- **C** 1.5
- **D** 2.0

Molar volume → 1mole = 22.4L

22.4L → 2g (H_2)

1.12L → ?

$$\frac{\cancel{1.12L} \times \cancel{2g}}{\cancel{22.4L}} = \boxed{0.1g}$$

$$\boxed{W = Z \times I \times t}$$

$$0.1 = \frac{\text{Eq wt of H}}{96{,}500} \times I \times 965$$

$$0.1 = \frac{1}{\cancel{96{,}500}} \times I \times \cancel{965}$$

$$0.1 = \frac{I}{100}$$

$$\boxed{I = 10A}$$

Eq. wt of Hydrogen

$$\frac{{}^{1}H}{H^{+1}} = \frac{1}{1}$$

= 1g

Question

$[H^+] = c\alpha$
$= 0.1 \times 0.01$

$[H^+] = \boxed{10^{-3}} = 0.001 \text{ mol dm}^{-3}$

$pH = -\log[H^+] = -\log 10^{-3}$

The resistance of $\underline{0.1}$ M weak acid HA in a conductivity cell is 2×10^3 Ohm. The cell constant of the cell is 0.78 cm^{-1} and Λ_m^∞ of acid HA is $390 \text{ S cm}^2 \text{ mol}^{-1}$. The pH of the solution is

A 3.3

B 4.2

C 5

D 3 ✓

$R = 2 \times 10^3$ ohm.

$\frac{l}{a} = 0.78 \text{ cm}^{-1}$

$\Lambda_m^\infty = 390 \text{ S cm}^2 \text{mol}^{-1}$

$pH = ?$

$c = 0.1$

$\boxed{\Lambda_m^c} = \frac{1000 \times K}{c}$

$= \frac{1000 \times 0.39 \times 10^{-3}}{0.1} = \frac{39 \times 10}{100} = \boxed{3.9}$

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

$= \frac{1}{2 \times 10^3} \times 0.78$

$\boxed{K = 0.39 \times 10^{-3} \text{ S cm}^{-1}}$

$\Lambda_m = 3.9 \text{ S cm}^2 \text{mol}^{-1}$

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

$= \frac{3.9}{390} = \frac{\cancel{39}}{10 \times \cancel{39} \times 10} \underline{\underline{0.01}}$

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

$[H^+] = c\alpha$

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

✓ $\boxed{\Lambda_m^c} = \frac{1000 \times K}{c}$

✓ $K = \frac{1}{R} \times l/a$

$[H^+] = 10^{-5}$

$\boxed{pH = 5}$

$[H^+] = 10^{-6}$

$\boxed{pH = 6}$

HCl

CH_3COOH

Acid $\longrightarrow H^{\oplus}$

Cx

$Cx\alpha \rightarrow$ No. of ions.

$\log a^b = b\log a$

$\log_{10} 10 = 1$

$H_2SO_4 \longrightarrow 2H^+ + SO_4^{2-}$

$C\ 10^{-3} = 2\times 10^{-3}$

$-\log 10^{-3} = -\left(-3\log_{10} 10\right)$

$= 3$

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

$pH = -\log[2\times 10^{-3}]$

$c\alpha$

$H_3PO_4 \longrightarrow 3H^{\oplus} + PO_4^{3-}$

$0.1 \longrightarrow 0.1\times 3$

$\longrightarrow 3\times 10^{-1}$

$[H^+] \rightarrow pH$

Question

Consider the following 4 electrodes

$Ag^+ + e^{\ominus} \longrightarrow Ag$

A : Ag^+ 10^{-4} (0.0001 M)/Ag(*s*) ;

B : Ag^+ 10^{-1} (0. 1 M)/Ag(*s*) ;

C : Ag^+ 10^{-2} (0.01 M)/Ag(*s*) ;

D : Ag^+ 10^{-3} (0.001 M)/Ag(*s*) ; $E^{\circ}_{Ag^+/Ag} = +0.80$ V

$$E_{cell} = E^{\circ}_{cell} - \frac{0.059}{n} \log \frac{[Ag]}{\boxed{[Ag^+]}}$$

$$E_{cell} = E^{\circ}_{cell} + \frac{0.059}{1} \log [Ag^+]$$

Then reduction potential ($\boxed{E_{cell}}$) in volts of the electrodes in the order.

~~A~~ B > C > D > A

B ~~C >~~ D > A > B

C A > D > ~~C >~~ B

D ~~A >~~ B > C > D

Question

Consider the following electrodes

$Zn^{2+} + 2e^{-} \longrightarrow Zn$.

P = Zn^{2+} (0.0001 M)/Zn, Q = Zn^{2+} (0.1 M)/Zn
R = Zn^{2+} (0. 01 M)/Zn, S = Zn^{2+} (0.001 M)/Zn
$E^\circ(Zn/Zn^{2+}) = -0.76$ V electrode potentials of the above electrodes in volts are in the order

A P > S > R > Q

B S > R > Q > P

C Q > R > S > P

D P > Q > R > S

Thank

You