Electrochemistry

Area of chemistry which is concerned with interconversion of chemical and electrical Energy.

Electric conduction :-

(A) Metallic conduction :- Conduction through metal involves a direct flow of electron from one point to the other

Outermost e^- are from conduction bond. and e^- are free to move.

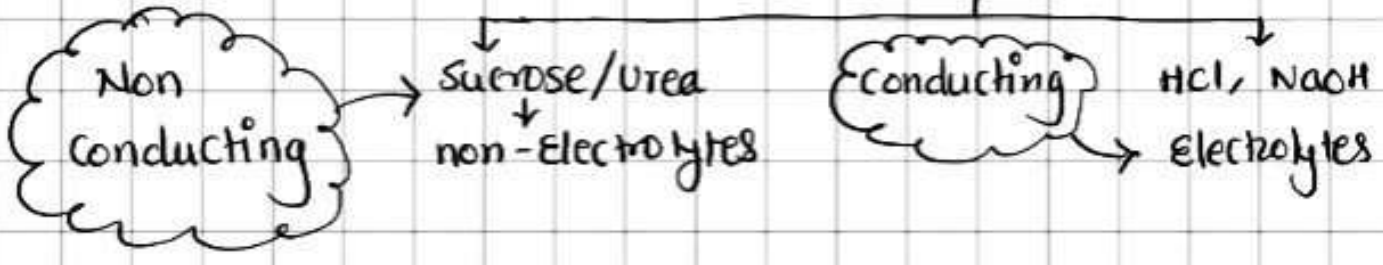
(B) Electrolytic or Ionic conduction :- conduction of Electric current by the movements of ion of the electrolyte.

The movement of ion through molten electrolytes or aq. solution of electrolytes

eg:- Ionic salts, Strong / weak electrolytes.

Information provided by measurement of conductivities of Solution

① Conducting or non conducting → by conductivity of sol^n
Nature of Solution



② On the basis of high or low Electrical Conductivity Electrolytes are classified as → Strong Electrolyte, Weak Electrolyte

Electrical conductance of the solution

According to the ohm's law

$V = IR$ $\quad \therefore R = V/I$ $\quad \Omega = \frac{\text{volt}}{\text{Ampere}}$

$\Omega = VA^{-1}$

V = potential difference

I = electric current

R = Resistance

G is Electric conductance

$\therefore G = 1/R$

SI unit of G Siemens (S)

$\therefore S = \Omega^{-1} = AV^{-1}$

$C = As \quad \therefore S = CV^{-1}s^{-1}$

$R \propto l/a$ (l: length, a: cross section)

$\boxed{R = \rho l/a}$ where ρ = resistivity

conductivity

$G = 1/R$ So $G \propto a/l$ $\quad G = K a/l$

K = conductivity

Conductivity of solution of Electrolyte is called Electrolytic conductivity

K is conductance of unit volume = $1cm^3$ or $1m^3$

$K = G\, l/a = 1/R\, l/a \quad \therefore K = 1/\rho$ ← resistivity

Molar conductivity

(Λ lambda)

$\boxed{\Lambda = K/C}$ where C = molar concentration

$\Lambda = \frac{Sm^{-1}}{mol\ m^{-3}}$

of units

$K = Sm^{-1}$

$C = mol\ m^{-3}$

$\therefore \Lambda = Sm^2mol^{-1}$

also $\Lambda = \Omega^{-1}cm^2mol^{-1}$

Significance of molar conductivity

Molar conductivity

→ Electrical conductance generated by all the ion in 1 mol of the Electrolyte

Relation between K & Λ

K = ele. conductance of $1cm^3$ of sol^n

Λ = Volume V in cm^3 containing 1 mol Electrolyte show Electric conductance Λ

$$\Lambda = KV$$

→ K is $1cm^3$ conductanc in total volume V

∴ total conductance is KV which is molar conductance

Concentration of solution

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

$$= \frac{C \text{ mol } L^{-1}}{1000 \text{ cm}^3 L^{-1}} = \frac{C}{1000} \text{ mol } cm^{-3}$$

$$V = \frac{1}{C} \quad \therefore V = \frac{1000}{C} \text{ cm}^3 \text{mol}^{-1}$$

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

Variation of conductivity with concentration

⇒ conductivity depends on no current carrying ions present in unit volume

⇒ on dil. no of ion increase & degree of dissociation

⇒ dil decrease con. so conductivity decrease

Variation of molar conductivity with concentration

⇒ molar conductivity is the electrical conductance of 1 mol of an Electrolyte in given volume

⇒ increase in no of moles will increase molar conductivity

Variation of molar conductivity with concentration

(a) Strong Electrolytes :- The molar conductivity of solution of Strong Electrolyte increase rapidly with dilution.

Molar conductivity at zero concentration OR infinite dilution
↓
(the limiting value 0·001M to 0·0001M)

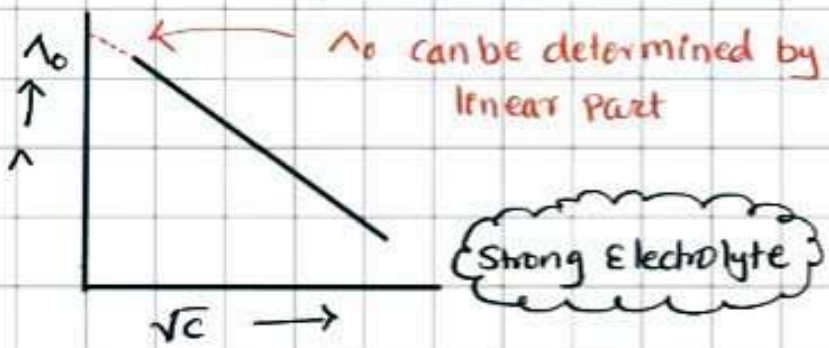
$[\Lambda_0]$

↳ means no change in conductivity on further dilution

⇒ Molar conductivity of strong electrolyte varies linearly with sq. root of concentration

$$\Lambda = \Lambda_0 - a\sqrt{C}$$

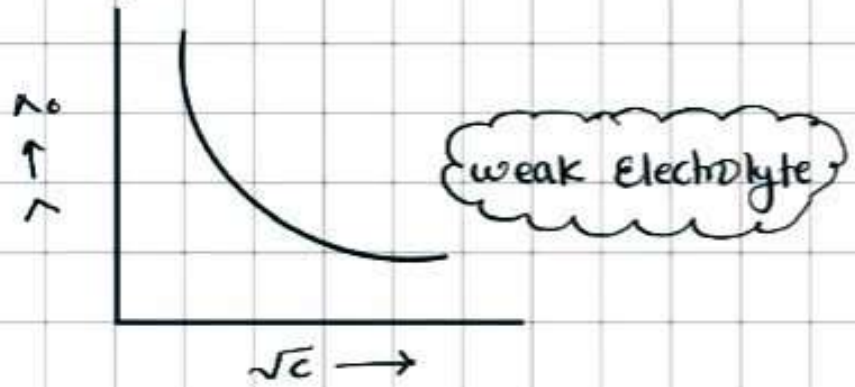
where : a ⇒ constant



(B) weak Electrolytes :- The molar conductivity of weak Electrolyte increase rapidly on dilution

⇒ The value of Λ is lower then Λ_0

Graph Λ vs $\sqrt{C}$ (concentration)



Kohlrausch law of Independent migration of ion

at infinite dilution each ion migrates independent of co-ion & contribute to total molar conductivity of electrolyte & not on nature of other ion.

$$\Lambda_o = n_\oplus \lambda^\circ_\oplus + n_\ominus \lambda^\circ_\ominus$$

$\lambda_\oplus$, $\lambda_\ominus$] Molar conductivity of cation & anion

$n_\oplus$, $n_\ominus$] No. of moles of cation & anion

Applications of Kohlrausch theory

⇒ to calculate molar conductivity of an Electrolyte at Zero Concentration

eg:- $\Lambda_o(KCl) = \lambda^\circ_{K^\oplus} + \lambda^\circ_{Cl^\ominus}$

$\Lambda_o[Ba(OH)_2] = \lambda^\circ_{Ba^{+2}} + 2\lambda^\circ_{OH^\ominus}$

⇒ Calculation of weak electrolytes from strong Electrolyte

eg:- $\Lambda^\circ(CH_3COOH)$

we have HCl, NaCl, CH_3COONa — strong Electrolytes

$\Lambda_o(HCl) + \Lambda_o(CH_3COONa) - \Lambda_o(NaCl)$

$= \lambda^\circ_{H^\oplus} + \lambda^\circ_{Cl^-} + \lambda^\circ_{CH_3COO^-} + \lambda^\circ_{Na^+} - \lambda^\circ_{Na^+} - \lambda^\circ_{Cl^-}$

HCl + CH_3COOH − NaCl

$= \lambda^\circ_{CH_3COO^-} + \lambda^\circ_{H^+}$

$= (\Lambda^\circ CH_3COOH)$

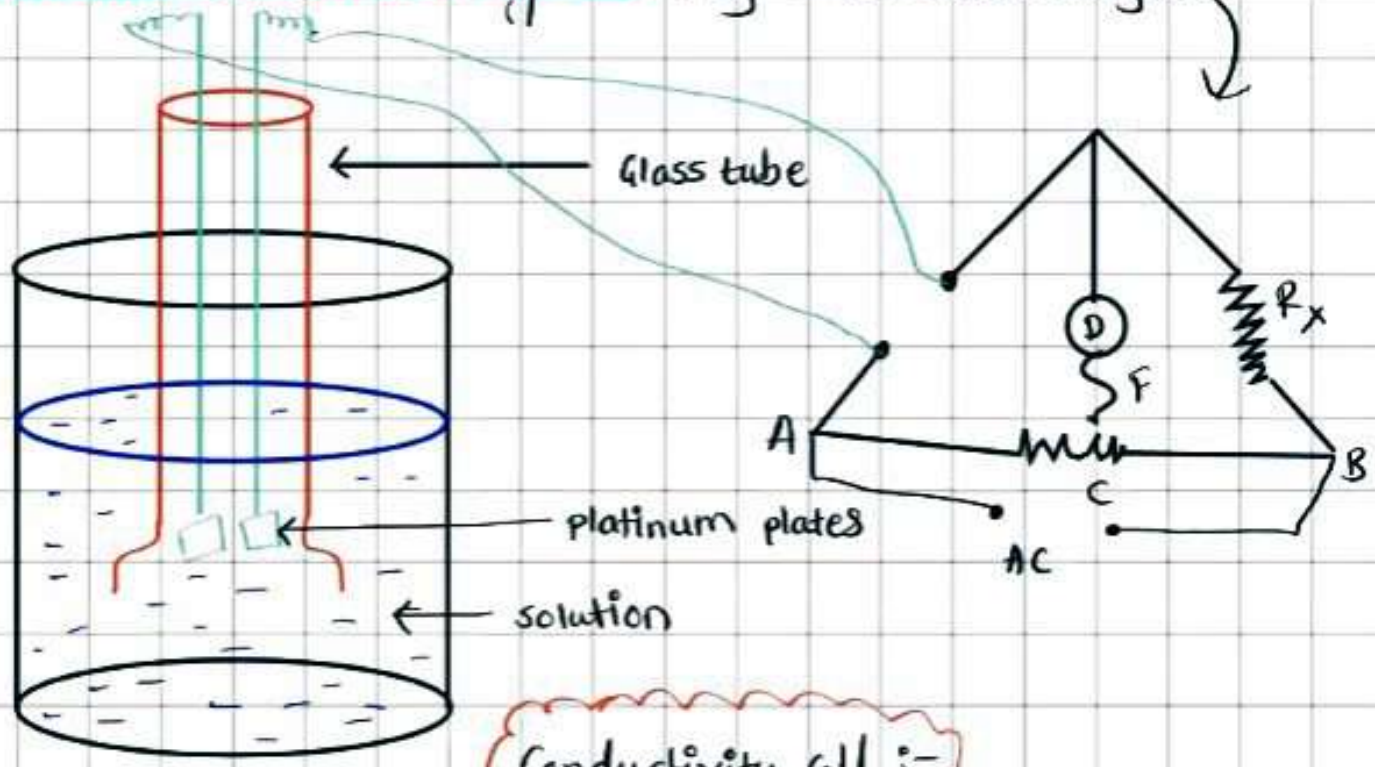
solve problems

Molar Conductivity and degree of dissociation of weak Electrolyte

$$\alpha = \frac{\Lambda_c}{\Lambda_o}$$

Λ_c = molar conductivity of weak electrolyte at conc. 'c'

Measurement of conductivity # (By wheatstone bridge)



Conductivity cell :-

Cell constant :-

$$K = 1/R \; 1/a$$

Cell constant $= 1/a$

SI unit m^{-1}/cm^{-1}

$$K = \frac{\text{cell constant}}{R}$$

→ Determination of cell constant ←

According to wheatstone bridge

$$\frac{R_{solution}}{l(AC)} = \frac{R_x}{l(BC)}$$

$$\therefore \quad R_{solution} = \frac{l(AC)}{l(BC)} \times R_x$$

Cell constant $= K_{KCl} \times R_{solution}$

② Determination of conductivity of given solution

$$K = \frac{\text{Cell constant}}{R \text{ solution}}$$

③ Calculation of molar conductivity

$$\Lambda = \frac{1000K}{C}$$

Electrochemical reaction

⇒ chemical reaction occuring in Electrochemical cell
⇒ involves transfer of electron
⇒ It is redox reaction

Electrodes # (A) cathode :- Electrode at which reduction takes place
(B) Anode :- Electrode at which oxidation takes place

Types of Electrochemical cell

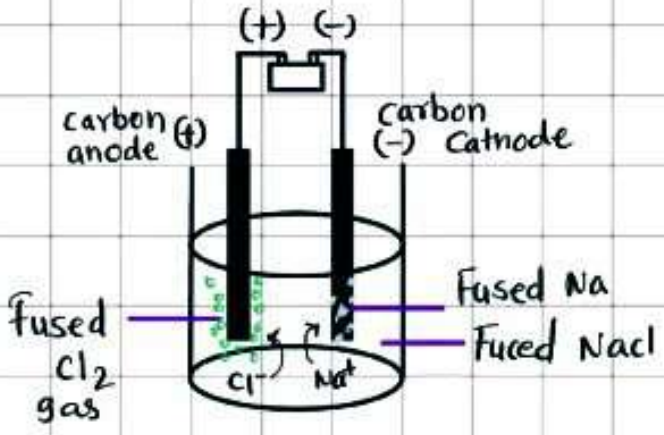
(A) Electrolytic cell :- Non-spontaneous reaction (Electrolysis)
⇒ convert Electrical Energy → chemical Energy.
⇒ Anode (+) cathode (-)

(B) Galvanic or voltaic cell :- Spontaneous chemical reaction produce Electricity.
⇒ Chemical Energy ⟶ Electrical Energy
⇒ Anode (-) cathode (+)

Electrolytic Cell

① Electrolysis of Molten NaCl

ⓐ Construction of cell –



ⓑ Cell Reaction :-

oxidation half reaction at anode

$$2Cl^-_{(l)} \longrightarrow Cl_{(g)} + Cl_{(g)} + 2e^-$$

$$Cl_{(g)}\ Cl_{(g)} \longrightarrow Cl_{2\,(g)}$$

$$2Cl^-_{(l)} \longrightarrow Cl_{2\,(g)} + 2e^- \quad \text{(oxidation)}$$

Reduction half at cathode :

$$Na^{\oplus}_{(l)} + e^{\ominus} \longrightarrow Na_{(l)}$$

Net cell reaction

$$2Cl^{\ominus} \longrightarrow Cl_2 + 2e^- \quad \text{(anode) oxidation}$$

$$2Na^{\oplus}_{(l)} + 2e^{\ominus} \longrightarrow 2Na_{(l)} \quad \text{(cathode) reduction}$$

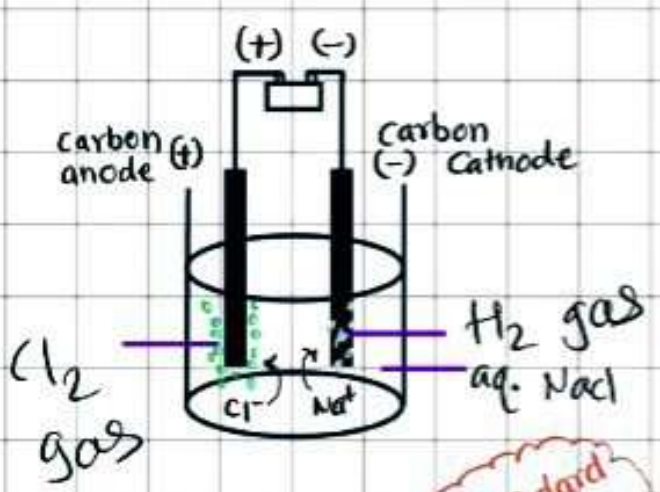
$$2Na^{\oplus} + 2Cl^{\ominus} \longrightarrow 2Na_{(l)} + Cl_{2\,(g)}$$

Results of Electrolysis of molten NaCl

(i) pale green Cl_2 gas released at anode

(ii) Molten silvery-white sodium is formed at cathode

* Electrolysis of aqueous NaCl *



Cell reaction :

Reduction half at cathode

i) $Na^+_{(aq)} + e^- \rightarrow Na$

ii) $2H_2O + 2e^- \rightarrow H_{2(g)} + OH^-_{(aq)}$

Standard Reduction potential → SRP of water is heighes then $Na^{\oplus}$

∴ there is red^n of H_2O at cathode

Oxidation half reaction at anode

i) $2Cl^-_{(aq)} \longrightarrow Cl_2 + 2e^-$ → $Cl^{\ominus}$ greater tendancy to get oxidised

ii) $2H_2O \longrightarrow O_2 + 4H^{\oplus} + 2e^-$

Overall cell reaction

$2Cl_{2(aq)} \longrightarrow Cl_{2(g)} + 2e^-$ (oxidation at anode)

$2H_2O_{(l)} + 2e^- \longrightarrow H_{2(g)} + Cl_{2(g)} + 2OH^-_{(aq)}$ (reduction at cathode)

$2Cl^-_{(aq)} + 2H_2O_{(l)} \longrightarrow Cl_{2(g)} + H_2 + OH^{\ominus}_{(aq)}$

Results of electrolysis of aq NaCl

(i) H_2 gas liberated at cathode

(ii) Cl_2 gas released at anode

(iii) $Na^{\oplus}$ ions and $OH^{\ominus}$ ions remain unreacted

∴ solution become alkaline ($NaOH$)

Quantitative aspects of Electrolysis

(i) Calcution of quantity of electricity passed

$$Q = I \times t$$

units → $C = AS$ → Seconds

Coulomb Ampere

(ii) Calculation of moles of Electrons passed

$$= \frac{Q\ (C)}{96500\ (C/mol\ e^-)}$$

(iii) calculation of moles of product formed

$$= \frac{Q}{96500} \times \text{mole ratio}$$

$$= \frac{I \times t}{9600} \times \text{mole ratio}$$

$$\text{mole ratio} = \frac{\text{moles of product formed}}{\text{moles of } e^- \text{ required}}$$

(iv) Calculation of mass of product

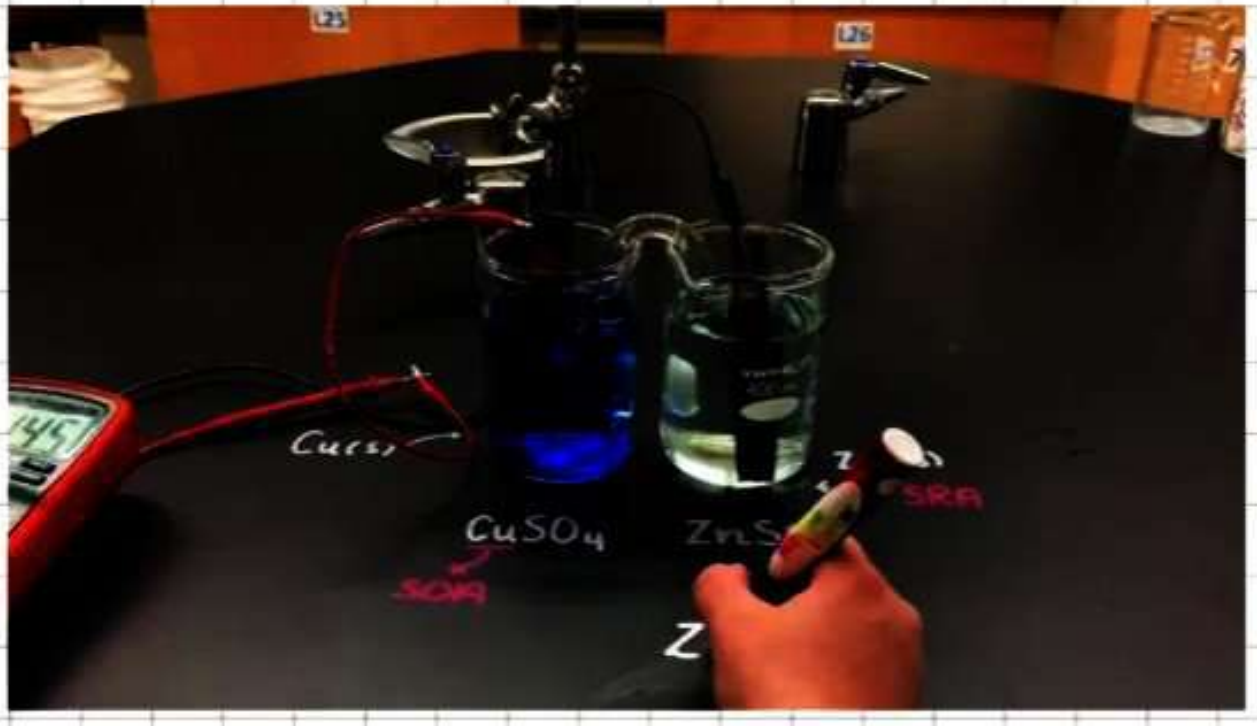
Mass of product (W) = Moles of product × Molar mass of product

$$\therefore = \frac{I \times t}{96500} \times \text{mole ratio} \times \text{Molar Mass of product}$$

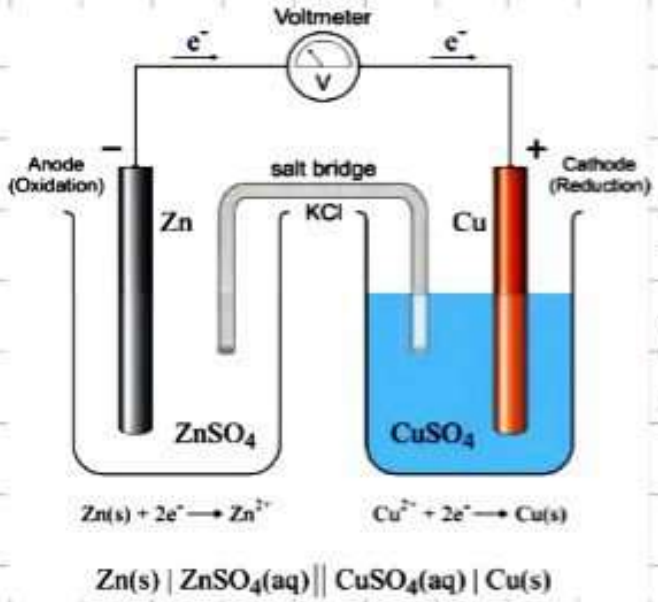
⇒ Electrolytes are connected in series

$$\frac{W_1}{(\text{mole ratio})_1 \times M_1} = \frac{W_2}{(\text{mole ratio})_2 \times M_2}$$

Galvanic or Voltaic Cell



Salt bridge:- It is a U-shape tube containing a saturated solution of an inert Electrolyte such as KCl or NH_4NO_3 & 5% agar solution



Function of the salt bridge

i) It provides an electrical contact bet^n two solution & complete the electrical circuit

ii) It prevents mixing of two solution

iii) It maintains electrical neutrality in both the solution by transfer of ion

Formation or short notation of galvanic cell

① The metal electrode or the inert electrodes are denoted by vertical lines placed at the end of the formula or short notation

← extreme left [Anode] (−) extreme Right → (cathode)(+)

② The insoluble species or gases are placed in the interior position adjesent to the metal electrode

③ The aq sol^{n} of ions placed at the middle of the cell Formula

④ A single vertical line between two phases indicates the phase boundary. It Indecates the direct contact between them

⑤ A double verticle line bet^{n} two sol^{n} indicates that they are conected by salt bridge

⑥ The additional information such as concentration, phases etc also given.

⑦ Single half cell is written in the order

→ aq sol^{n} of Rn 1st then solide electrode

eg:- $Zn^{2\oplus}_{(1M)} \mid Zn_{(s)}$

eg:- of short notation of galvanic cell

① two half cells are a) $Mg^{+2}_{(1M)} \mid Mg_{(s)}$ b) $Ag^{+}_{(1M)} \mid Ag_{(s)}$

Cell representation

$$Mg_{(s)} \mid Mg^{2+}_{(1M)} \,||\, Ag^{+}_{(1M)} \mid Ag_{(s)}$$

Anode — salt bridge — Cathode

writing of cell reaction

↳ best with examples

Electrode potential & Cell potential

Electrode potential - potential difference established due to electrode half reaction occuring at the electrode surface, is the electrode potential.

overall cell potential (Emf) (Electromotive Force)
the contribution of each electrode OR
algebric sum of the electrode potential

$$E_{cell} = E_{(oxi)}^{anode} + E_{(Red)}^{Cathode}$$

Standard potential [E°]
Standard condition 1 M concentration
1 atm pressure

$$E^\circ_{cell} = \underset{(anode)}{E^\circ_{oxi}} + \underset{(cathode)}{E^\circ_{red}}$$

According to the IUPAC Standard Electrode potential is taken as Standard Reduction potential [SRP]

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

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

This Formula is used To calculate cell potential [EMF] of standard cell.

Dependence of cell potential on concentration

(Nernst equation)

$$aA + bB \longrightarrow cC + dD$$

The cell voltage is given by

$$E_{cell} = E^{\circ}_{cell} - \frac{RT}{nf} \ln \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

$$= E^{\circ}_{cell} - \frac{2.303\,RT}{nf} \log_{10} \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

at 25°C

$$\frac{2.303\,RT}{F} = 0.0592$$

F = Faraday = 96500 C

R = gas const. = $8.314\ J\,K^{-1}\,mol^{-1}$

T = temp in 'K' kelvin

n = no. of moles of e^-

$$\therefore E_{cell} = E^{\circ}_{cell} - \frac{0.0592}{n} \log_{10} \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

If con is not Standard then emf of cell can be calculated by nernst equation

If con. of reactant and product is (1 M) then, $E_{cell} = E^{\circ}_{cell}$

(i) Calculation of cell potential

$$Cd_{(s)} \mid Cd^{+2}_{(aq)} \parallel Cu^{2\oplus}_{(aq)} \mid Cu$$

Cell reaction

$Cd_{(s)} \longrightarrow Cd^{2+} + 2e$ (oxidⁿ) at anode

$Cu^{+2}_{(aq)} + 2e \longrightarrow Cu$ (Redⁿ) at cathode

Net cell Reaction

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

$n = 2$

$\therefore$

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0592}{2} \log_{10} \frac{[Cd^{+2}]}{[Cu^{+2}]}$$

Concentration of solid & pure liq. are taken to be unity (1)

(ii) calculation of Electrode potential

$$Zn^{+2}_{(aq)} \mid Zn_{(s)}$$

$Zn^{+2}_{aq} + 2e^{-} \longrightarrow Zn_{(s)}$ reduction

$$E_{Zn} = E^{\circ}_{Zn} - \frac{0.0592}{2} \log_{10} \frac{1}{[Zn^{2+}]}$$

$$= E^{\circ}_{Zn} + \frac{0.0592}{2} \log_{10} [Zn^{+2}]$$

Thermodynamics of galvanic cell

Gibbs energy of cell reaction and cell potential

Electrical work = amount of charge passed × cell potential [E_{cell}]

charge passed = nF ∴ = nf E_{cell}

But

Electrical energy = $-\Delta G$

∴ $-\Delta G = nf\, E_{cell}$

$\Delta G = -nf\, E_{cell}$

at standard $\Delta G^\circ = -nF E^\circ_{cell}$

ΔG° → extensive

E°_{cell} → intensive

Standard cell potential & equilibrium constant

$\Delta G = -nf\, E^\circ_{cell}$ But $\Delta G = -RT \ln K$

∴ $-nf E^\circ_{cell} = -RT \ln K$

$nf\, E^\circ_{cell} = RT \ln K$

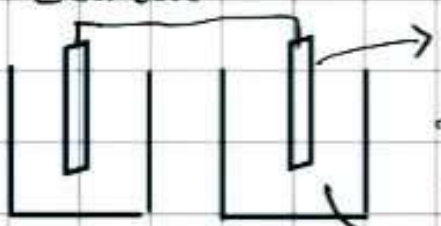
$E^\circ_{cell} = \frac{RT}{nF} \ln K$

$= \frac{2.303RT}{nF} \log_{10} K$ But $\frac{2.303RT}{F} = 0.0592$

$$\therefore E^\circ_{cell} = \frac{0.0592}{n} \log_{10} K$$

Reference Electrode

it is used to calculate the potential of the Electrode



→ one of them consider as zero

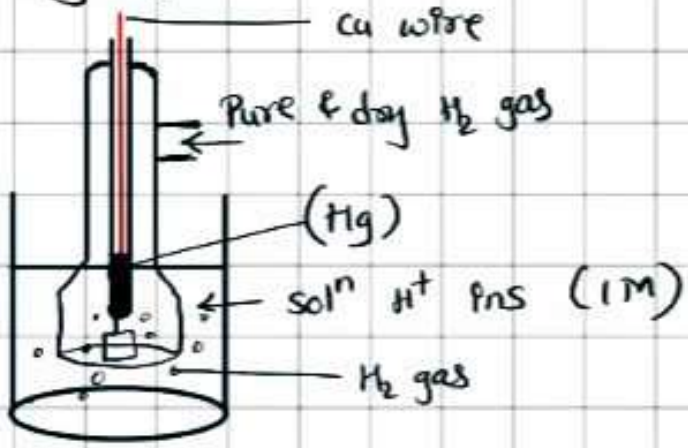
then E_{cell} will give value of another Electrode

∴ → This type of electrodes are known as reference electrode

eg:-

(1) Standard Hydrogen Electrode (SHE)

Construction



Formulation $H^+_{(1m)} \mid H_{2\,(g)}(1atm) \mid Pt$

Electrode reaction

$$2H^+(1M) + 2e^- \longrightarrow H_2\ (g, 1atm)$$

$E^\circ_{H_2} = 0.000\ V$

Application of SHE

SHE is used as reference electrode

For eg:- determine the standard potential of

$Zn^{+2}_{(1M)} \mid Zn_{(s)}$

combine it with SHE

$Zn \mid Zn^{+2}_{(1M)} \parallel H^+_{(1M)} \mid H_{2\,(g)}(1atm) \mid Pt$

$E^\circ_{cell} = E^\circ_{H_2} - E^\circ_{Zn}$ But $E^\circ_{H_2} = 0$ ∴ $\boxed{E^\circ_{cell} = -E^\circ_{Zn}}$

Difficulties

① difficult to obtain pure & dry Hydrogen
② pressure cannot be maintain 1 atm
③ con. of sol^n cannot be exactly 1 M

Hydrogen gas Electrode

$H^{\oplus}_{aq} \mid H_2(g, P_{H_2}) \mid Pt\ [H^+]$

P_{H_2} → Pressure of Hydrogen gas P_H not (1) which means it is not standard

Electrode reaction

$$2H^{\oplus}_{(aq)} + 2e^- \longrightarrow H_2\ (g, P_{H_2})$$

$$E_{H_2} = E^\circ_{H_2} - \frac{0.0592}{2} \log_{10} \frac{P_{H_2}}{[H^+]^2}$$

$$= 0 - \frac{0.0592}{2} \log_{10} \frac{P_{H_2}}{[H^+]^2}$$

$E^\circ_{H_2} = 0\ V$

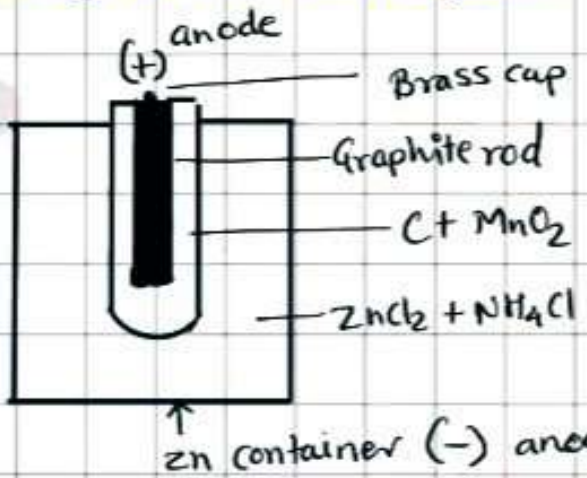
Galvanic cells useful in day to day life

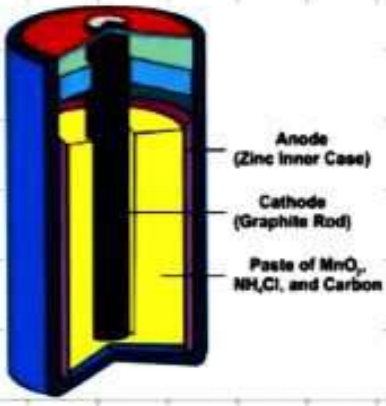
Types (A) primary voltic cell :- cannot be recharged
eg- dry cell

(B) Secondary voltic cell :- can be recharged
while recharge they will act as Electrolytic cell

① Dry cell (leclanche' cell)

Construction





Cell reaction

Ⓐ oxidation at anode

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

Ⓑ Reaction at cathode

① $2\,NH_4^{\oplus}{}_{(aq)} + 2e^- \longrightarrow 2\,NH_3{}_{(aq)} + H_2{}_{(g)}$

② $H_2{}_{(g)} + 2\,MnO_2{}_{(s)} \longrightarrow Mn_2O_3{}_{(s)} + H_2O_{(l)}$

$$2\,NH_4^{\oplus}{}_{(aq)} + 2\,MnO_2{}_{(s)} + 2e^- \longrightarrow Mn_2O_3 + 2NH_3{}_{(aq)} + H_2O_{(l)}$$

Ⓒ Net cell reaction

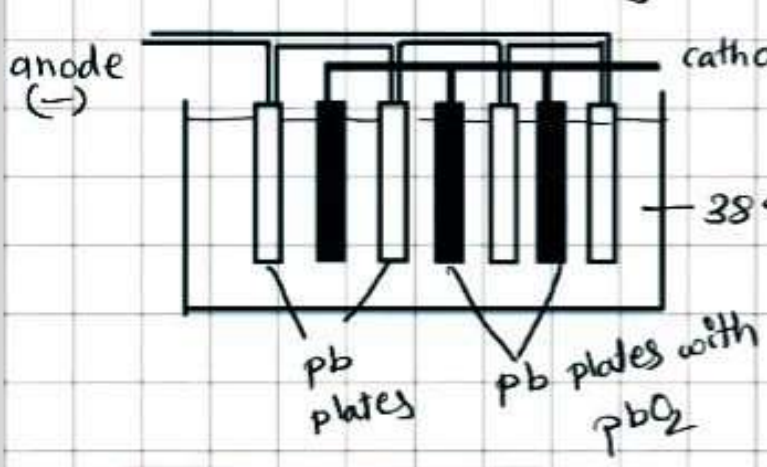
$$Zn_{(s)} + 2NH_4^{\oplus} + 2MnO_2 \longrightarrow Zn^{+2} + Mn_2O_3 + 2NH_3 + H_2O$$

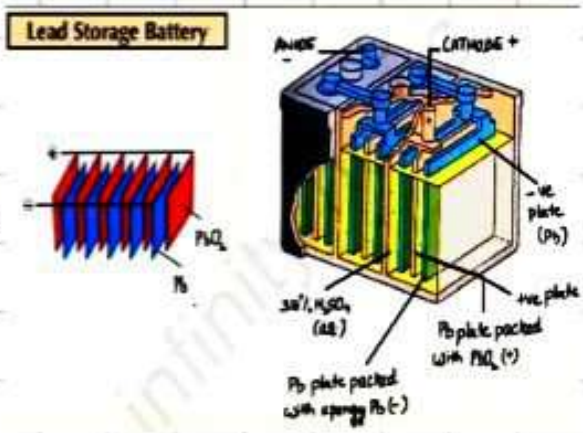
$$Zn^{+2} + 4\,NH_3{}_{(aq)} \longrightarrow [Zn(NH_3)_4]^{+2}{}_{(aq)}$$

use of dry cell :-

Source of power in flashlight, portable radio tape recorders etc.

Lead storage battery (lead accumulator)





Notation of the Cell

$$Pb_{(s)} \mid PbSO_{4\,(s)} \mid 38\%\ H_2SO_{4\,(aq)} \mid PbSO_{4\,(s)} \mid PbO_{2\,(s)} \mid Pb_{(s)}$$

(A) Cell reaction during discharge

a) oxidation at anode

$$Pb_{(s)} \longrightarrow Pb^{2\oplus}_{(aq)} + 2e^{\ominus} \quad (\text{oxid}^{n})$$

$$Pb^{+2}_{aq} + SO_4^{2\ominus} \longrightarrow PbSO_{4\,(s)} \quad (\text{ppt})$$

$$Pb_{(s)} + SO_{4\,(aq)}^{2\ominus} \longrightarrow PbSO_{4(s)} + 2e^{\ominus}$$

b) Reduction at cathode

$$PbO_{2(s)} + 4H^{+}_{(aq)} + 2e^{\ominus} \longrightarrow Pb^{2\oplus} + H_2O$$

$$Pb_{(s)} + SO_{4\,(aq)}^{2\ominus} \longrightarrow PbSO_{4\,(s)}$$

$$PbO_{2\,(s)} + 4H^{\oplus}_{(aq)} + SO_4^{2\ominus} + 2e^{\ominus} \longrightarrow PbSO_4 + 2H_2O$$

Net cell reaction

$$Pb_{(s)} + PbO_{2\,(s)} + 4H^{\oplus}_{(aq)} + 2SO_4^{2\ominus} \longrightarrow 2PbSO_{4(s)} + 2H_2O$$

(B) Cell Reaction during recharge

During recharge cell function as Electrolytic cell

(a) oxidation at anode (+)

$$PbSO_{4(s)} + 2H_2O_{(l)} \longrightarrow PbO_{2(s)} + 4H^{\oplus}_{(aq)} + SO_{4\,(aq)}^{2\ominus} + 2e^{\ominus}$$

(b) Reduction at cathode (−)

$$PbSO_{4(s)} + 2e^{-} \longrightarrow Pb_{(s)} + SO_{4\,(aq)}^{2\ominus}$$

(c) Net cell reaction

$$2PbSO_{4(s)} + 2H_2O_{(l)} \longrightarrow Pb_{(s)} + PbO_{2(s)} + 2H_2SO_{4(aq)}$$

Applications of lead accumulator

1) used as source of direct current in the laboratory
2) a 12 V lead storage battery constructed by connecting six 2 V cells in series is used in automobiles & inverters

Nickel-Cadmium or NICAD storage cell

⇒ it is secondary cell ⇒ cathode is Ni(IV) oxide
Anode is Cd-metal

Cell Reaction

$$Cd_{(s)} + 2OH^-_{(aq)} \longrightarrow Cd(OH)_{2(s)} + 2e^{\ominus}$$

$$NiO_{2(s)} + 2H_2O_{(l)} + 2e^{\ominus} \longrightarrow Ni(OH)_2 + 2OH^{\ominus}$$

$$Cd_{(s)} + NiO_{2(s)} + 2H_2O_{(l)} \longrightarrow Cd(OH)_2 + Ni(OH)_2$$

uses: - longer life than other dry cell
used in watches, calculators etc

Mercury battery

⇒ secondary cell ⇒ zinc (Zn) anode
Carbon + Hg cathode

Cell Reaction

(anode oxid^n) $Zn(Hg) + 2OH^{\ominus}_{(aq)} \longrightarrow ZnO_{(s)} + H_2O_{(l)} + 2e^-$

(cathode Red^n) $HgO + H_2O + 2e^- \longrightarrow Hg_{(l)} + 2OH^{\ominus}_{(aq)}$

$$Zn(Hg) + HgO_{(s)} \longrightarrow ZnO_{(s)} + Hg_{(l)}$$

Fuel Cell

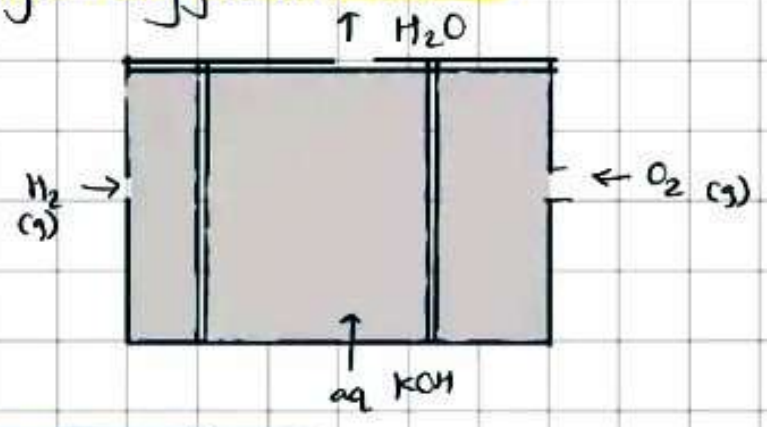
Function is based on combustion reaction are redox type & can be used to generate electricity.

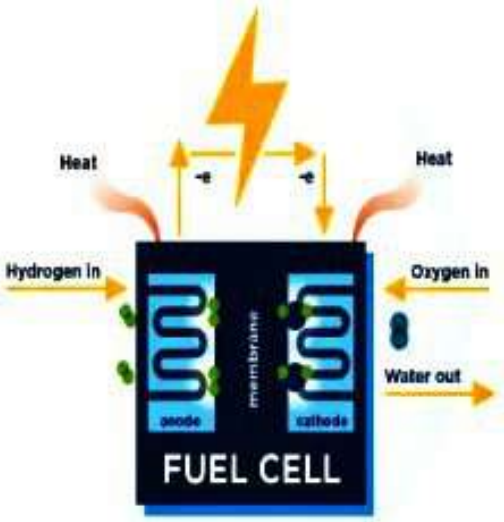
⇒ one of the reactant is fuel & other is reactant such as O_2, is oxidant

eg:- Hydrogen-oxygen fuel cell

⇒ Fuel is hydrogen (H_2) gas

Hydrogen-oxygen fuel cell





Cell Reaction

(i) oxidation at anode (−)

$$2H_{2(g)} + 4OH^{\ominus}_{(aq)} \longrightarrow 4H_2O_{(l)} + 4e^{\ominus}$$

(ii) Reduction at cathode (+)

$$O_{2(g)} + 2H_2O_{(aq)} + 4e^{\ominus} \longrightarrow OH^{\ominus}_{(aq)}$$

(iii) Net cell reaction

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)}$$

Advantages of fuel cell

(1) they are nonpolluting as the only reaction product is H_2O

(2) provides Electricity with High efficiency (70%)

(3) the reacting substance is continuously supplied ∴ not have to be discarded on consuming of chemicals

Drawbaks of fuel cell

H_2 gas is hazardous & preparation cost is high

Applications :- (1) experimental basis in automobiles
(2) For electrical power in space programe

Electrochemical series

Sr. No.	Electrode	Reduction Half reaction Oxidising agent → Reducing agent	E^0 (volts) At 25
1	$F^- \mid F_2 \mid Pt$	$F_2 + 2e^- \rightarrow 2F^-$	+2.87
2	$Au^+ \mid Au$	$Au^+ + e^- \rightarrow Au$	+1.68
3	$Ce^{4+}, Ce^{3+} \mid Pt$	$Ce^{4+} + e^- \rightarrow Ce^{3+}$	+1.61
4	$Au^{3+} \mid Au$	$Au^{3+} + 3e^- \rightarrow Au$	+1.50
5	$Cl^- \mid Cl_2 \mid Pt$	$Cl_2 + 2e^- \rightarrow 2Cl^-$	+1.36
6	$Pt^{2+} \mid Pt$	$Pt^{2+} + 2e^- \rightarrow Pt$	+1.20
7	$Br^- \mid Br_2 \mid Pt$	$Br_2 + 2e^- \rightarrow 2Br^-$	+1.08
8	$Hg^{2+} \mid Hg$	$Hg^{2+} + 2e^- \rightarrow Hg$	+0.854
9	$Ag^+ \mid Ag$	$Ag^+ + e^- \rightarrow Ag$	+0.799
10	$Hg_2^{2+} \mid Hg$	$Hg_2^{2+} + 2e^- \rightarrow Hg_2$	+0.790
11	$Fe^{3+}, Fe^{2+} \mid Pt$	$Fe^{3+} + e^- \rightarrow Fe^{2+}$	+0.771
12	$I^- \mid I_{2(s)} \mid Pt$	$I_2 + 2e^- \rightarrow 2I^-$	+0.535
13	$Cu^{2+} \mid Cu$	$Cu^{2+} + 2e^- \rightarrow Cu$	+0.337
14	$Pt \mid Hg \mid Hg_2Cl_2 \mid Cl^-$	$Hg_2Cl_2 + 2e^- \rightarrow 2Hg + 2Cl^-$	+0.242
15	$Ag \mid AgCl_{(s)} \mid Cl^-$	$AgCl + e^- \rightarrow Ag + Cl^-$	+0.222
16	$Cu^{2+} \mid Cu^+$	$Cu^{2+} + e^- \rightarrow Cu^+$	+0.153
17	$Sn^{4+}, Sn^{2+} \mid Pt$	$Sn^{4+} + 2e^- \rightarrow Sn^{2+}$	+0.15
18	$H^+ \mid H_2 \mid Pt$	$2H^+ + 2e^- \rightarrow H_{2(g)}$	0.0 (Definition)
19	$Pb^{2+} \mid Pb$	$Pb^{2+} + 2e^- \rightarrow Pb$	-0.126
20	$Sn^{2+} \mid Sn$	$Sn^{2+} + 2e^- \rightarrow Sn$	-0.136
21	$Ni^{2+} \mid Ni$	$Ni^{2+} + 2e^- \rightarrow Ni$	-0.257
22	$Co^{2+} \mid Co$	$Co^{2+} + 2e^- \rightarrow Co$	-0.280
23	$Cd^{2+} \mid Cd$	$Cd^{2+} + 2e^- \rightarrow Cd$	-0.403
24	$Fe^{2+} \mid Fe$	$Fe^{2+} + 2e^- \rightarrow Fe$	-0.440