

Unit - 3

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Electrochemistry

Electrochemistry → The branch of a science which deals with the production of electricity from chemical changes or chemical rxn and the use of electricity to bring out a chemical rxn.

Electrolysis → It is the process in which chemical change take place by the passage of electric current or electricity.

Electrolysis is carried out in a cell
i.e. electrolytic cell.

Electrolysis of molten ^{NaCl} cell →
 $\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^-$

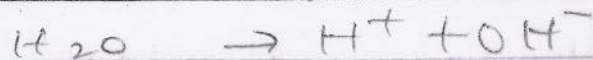
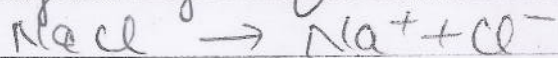
At cathode.



At anode $\text{Cl}^- - e^- \rightarrow \text{Cl}$



Electrolysis of aqueous NaCl →



At cathode $\text{H}^+ + e^- \rightarrow \text{H}$



At anode



Faraday's law of electrolysis →

Foraday's 1st law of electrolysis → According to this law the amount of any substance deposited or liberated at any electrode is directly proportional to the quantity of electricity passed through solution. @

If 'w' be the amount of substance deposited or liberated and 'Q' be the quantity of electricity passed then according to Foraday's first law

$$w \propto Q$$
$$w = ZQ \quad \left[\begin{array}{l} \text{where } Z \text{ is a constant} \\ \text{it is electrochemical} \\ \text{equivalent} \end{array} \right]$$

If 'I' be the current in ampere passed for time t then quantity of electricity

$$Q = I \times t$$

$$\therefore w = Z \times I \times t$$

If $I = 1$ Amp & $t = 1$ sec then $w = Z$

∴ Electrochemical equivalent is equal to the amount of substance deposited or liberated when 1 amp current is passed for 1 second.

Foraday's 2nd law of electrolysis → According to this law when same quantity of electricity is passed through different electrolytic solutions connected in series. The weight of substance deposited at electrode is directly proportional to their equivalent weight.

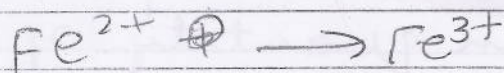
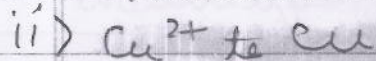
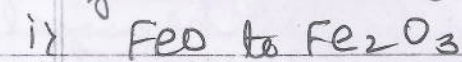
For eg. → let us consider two solutions $CuSO_4$ & $AgNO_3$ when same amount of electricity is passed through them connected in series

$$\frac{\text{then weight of Cu deposited}}{\text{wt. of Ag deposited}} = \frac{\text{eq. wt. of Cu}}{\text{eq. wt. of Ag}}$$

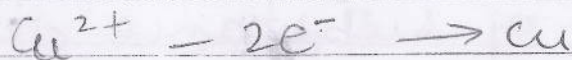
This law can also be express in terms of moles of e^- exchange during redox rxn.

Acc. to this law the mole of a substance deposited or liberated at an electrode is proportional to the no. of moles of e^- exchange during oxidation and reduction at that electrode.

eg \Rightarrow How many coulomb of electricity are required for the oxidation of one mole of



For oxidation of one mole of FeO to Fe_2O_3
electricity required = $1F$ i.e. $96500C$

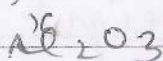


For oxidation of one mole of Cu^{2+} to Cu
electricity required = $2F$

$$= 2 \times 96500C$$

$$= 193000C$$

eg \Rightarrow How many Charge is required to produce 40g of Al from Al_2O_3

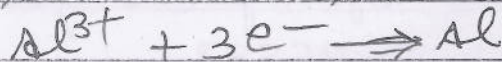


$$2x + (-2)(3) = 0$$

$$2x - 6 = 0$$

$$2x = 6$$

$$x = \frac{6}{2} = +3$$



For depositing 1 mole of Al charge required
= 3F

$$= 96500 \times 3$$

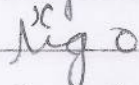
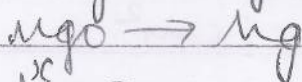
For depositing 27g of Al charge required
= 3 × 96500

For depositing 1g of Al charge required
= $\frac{3 \times 96500}{27}$

For depositing 10g of Al charge required
= $\frac{3 \times 96500}{27} \times 10$

$$= \frac{965000}{9} = 107222.3$$

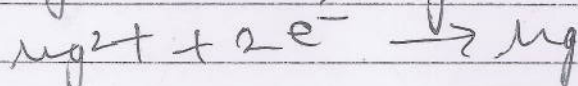
Q.10) How many charge is required for depositing 12g of Mg from MgO.



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

$$x - 2 = 0$$

$$x = +2$$



For one mole depositing 1 mole of Mg
Charge required = $2F$
= 2×96500

For depositing 24 g of Mg charge required
= 2×96500

For depositing 1g of Mg charge required
= $\frac{2 \times 96500}{24}$

For depositing 12g of Mg charge required
= $\frac{2 \times 96500}{24} \times \frac{12}{24}$
= 96500

Resistance \rightarrow It is the property of a substance which measures the obstruction to the flow of current.
units \rightarrow Ohm

Resistance of a conductor is directly proportional to its length and inversely proportional to its area.
Mathematically

$$R \propto \frac{l}{a}$$

$$R = \rho \frac{l}{a} \left[\begin{array}{l} \text{where } \rho \text{ is } \rho \text{ (or specific)} \\ \text{resistance or} \\ \text{resistivity} \end{array} \right]$$

Let $l = 1\text{cm}$ & $a = 1\text{cm}^2$
 $R = \rho$

Specific resistance or resistivity is the resistance of a substance having unit length (1cm) and unit area (1cm^2)

units $R = \rho \frac{l}{a}$

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

$$\rho = \text{ohm} \frac{\text{cm}^2}{\text{cm}}$$

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

Conductance $\rightarrow (C)$ It is the ease with which current can flow through a substance.

Conductance is inverse of resistance.

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

unit $\rightarrow \text{ohm}^{-1} = \text{mho}$

Specific conductance \rightarrow The conductance of a substance having unit length and unit area of cross section is called specific conductance and conductivity.

Specific conductance or conductivity is inverse of specific resistance or resistivity.

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

$$K = \frac{1}{\text{ohm cm}}$$

units $K = \text{ohm}^{-1} \text{cm}^{-1}$

Equivalent conductance \rightarrow It is the conducting power of all the ions produced

by dissolving 1 gm equivalent of electrolyte in the solution.

Mathematically

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

where C is the concentration of solution express in gm equivalent per litre -
units of C \rightarrow gm equivalent l^{-1}

$$\lambda_e = \frac{\text{ohm}^{-1} \text{cm}^{-1}}{\text{gm equivalent } l^{-1}}$$

$$\lambda_e = \frac{\text{ohm}^{-1} \text{cm}^{-1}}{\text{gm equivalent} / \text{cm}^3}$$

$$\lambda_e = \frac{\text{ohm}^{-1} \text{cm}^{-1} \text{cm}^3}{\text{gm equivalent}}$$

$$\lambda_e = \frac{\text{ohm}^{-1} \text{cm}^2}{\text{gm equivalent}}$$

$$\lambda_e = \frac{\text{ohm}^{-1} \text{cm}^2 \text{gm} (\text{equivalent})^{-1}}{(s)}$$

molar conductance \rightarrow It is the conductivity power of all the ions produce by dissolving one mole of electrolyte in the solution

Mathematically

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

where M is the concentration express in mole per litre

units of M = mole per litre

$$\lambda_m = \frac{\text{ohm}^{-1} \text{cm}^{-1}}{\text{mol } l^{-1}}$$

$$\lambda_m = \frac{\rho \text{hm}^{-1} \text{cm}^{-1} \text{cm}^3}{\text{mole}}$$

$$\lambda_m = \rho \text{hm}^{-1} \text{cm}^2 (\text{mole})^{-1}$$

Factors on which conductance or molar conductance depends → Conductance and molar conductance depends upon following factors

- 1) Nature of electrolyte → The conductance or molar conductance depends upon the nature of a electrolyte. If the electrolyte is strong then more be its conductance or molar conductance. If the electrolyte is weak then less be its conductance or molar conductance.
- 2) Concentration of solution → Strong electrolyte → The variation of molar conductance with concentration for a strong electrolyte can be explained by an equation

$$\lambda_m = \lambda_m^\infty - b\sqrt{C}$$

where λ_m^∞ = molar conductance

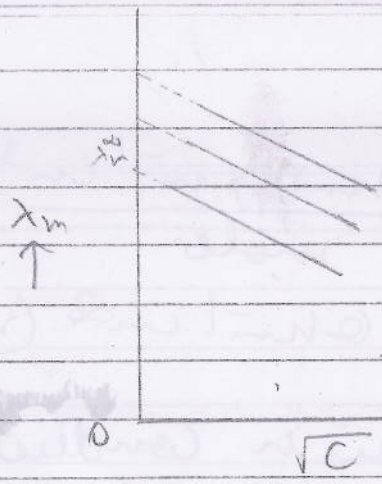
λ_m^∞ = molar conductance at infinite dilution or zero concentration.

b = constant

C = Concentration

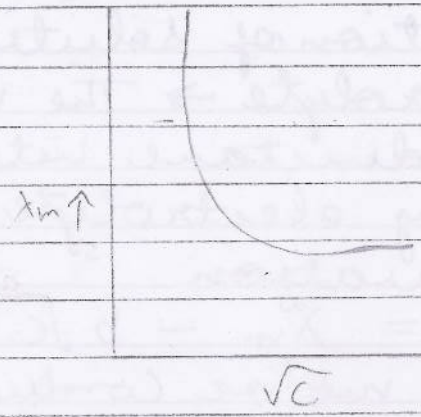
This equation is known as Debye Hückel Onsager equation.

The variation of molar conductance with conc. can be expressed by a graph b/w molar conductance & concentration.



From this graph we can calculate molar conductance at infinite dilution and for strong electrolyte by extra plotting a graph.

Weak electrolyte \rightarrow The variation of molar conductance and with concentration for a weak electrolyte can be explain by plotting a graph b/w molar conductance and concentration.



Kohlrausch's law \rightarrow Acc. to this law at infinite dilution or zero concentration each ion of the electrolyte makes a definite contribution towards the molar conductance of the electrolyte irrespective of the nature of other ion

with which it is attached.

Mathematically

$$\lambda_m^\infty = \nu_{(+)} \lambda_{m(+)}^\infty + \nu_{(-)} \lambda_{m(-)}^\infty$$

where λ_m^∞ = molar conductance at infinite dil

$\nu_{(+)}$ = no. of +ve ions

$\nu_{(-)}$ = no. of -ve ion

$\lambda_{m(+)}^\infty$ = molar conductance of +ve ions at infinite dilution

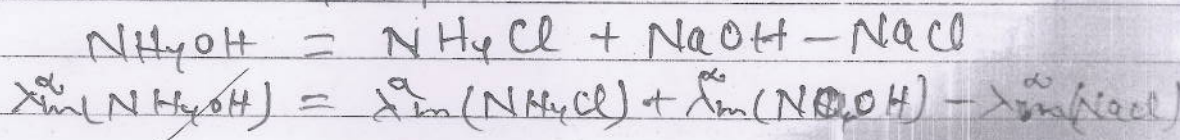
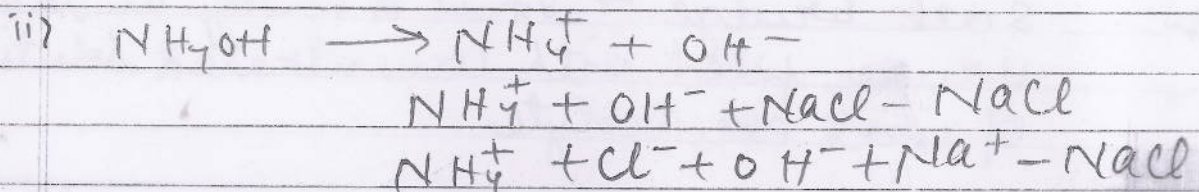
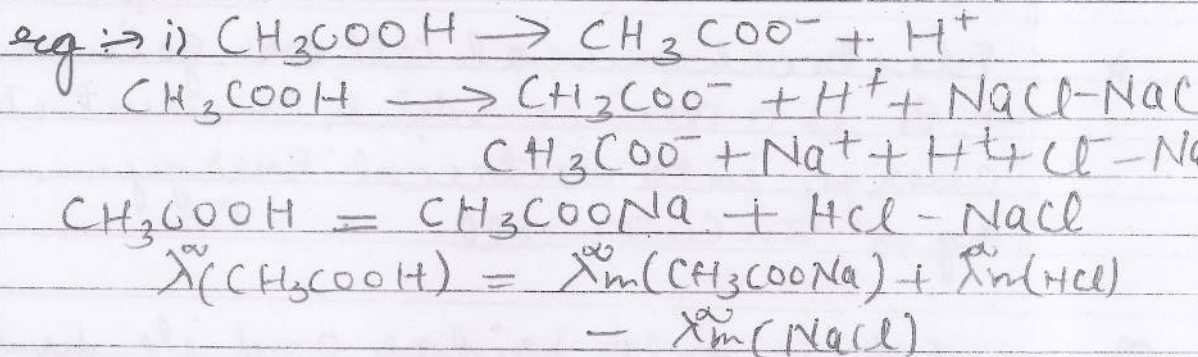
$\lambda_{m(-)}^\infty$ = molar conductance of -ve ion at infinite dilution

ing

$$\text{eg} \Rightarrow \lambda_m^\infty(\text{CH}_3\text{COOH}) = \lambda_m^\infty(\text{CH}_3\text{COO}^-) + \lambda_m^\infty(\text{H}^+)$$

22 Applications of Kohlrausch law \rightarrow

- 1) calculation of molar conductance of weak electrolyte \rightarrow molar conductance of weak electrolyte can be determine with the molar conductance of strong electrolyte at infinite dilution using Kohlrausch law.



eg \Rightarrow The molar conductance at infinite dilution of CH_3COONa , HCl & NaCl are 91, 425.9, 126 $\text{ohm}^{-1}\text{cm}^2\text{mol}^{-1}$. Calculate the molar conductivity of CH_3COOH at infinite dilution.

Sol: \Rightarrow

$$\lambda^\infty(\text{CH}_3\text{COOH}) = \lambda^\infty_{\text{H}^+} + \lambda^\infty_{\text{Cl}^-} - \lambda^\infty_{\text{Na}^+} - \lambda^\infty_{\text{Cl}^-}$$

$$\lambda^\infty(\text{CH}_3\text{COOH}) = \lambda^\infty(\text{CH}_3\text{COONa}) + \lambda^\infty(\text{HCl}) - \lambda^\infty(\text{NaCl})$$

$$\lambda^\infty(\text{CH}_3\text{COONa}) = 91$$

$$\lambda^\infty(\text{HCl}) = 425.9$$

$$\lambda^\infty(\text{NaCl}) = 126$$

$$\lambda^\infty(\text{CH}_3\text{COOH}) = 91 + 425.9 - 126$$

$$= 390.95 \text{ cm}^2 \text{ mol}^{-1}$$

2) Calculation of degree of dissociation \rightarrow
The degree of dissociation of weak electrolyte is represented as $\alpha = \frac{\lambda_m}{\lambda_m^\infty}$

λ_m^∞ can be calculated with the help of Kohlrausch law and thus degree of dissociation α can be calculated.

where $\lambda_m^\infty =$ molar conductance at any conc. C
 $\lambda_m^\infty =$ molar conductance at infinite dilution.

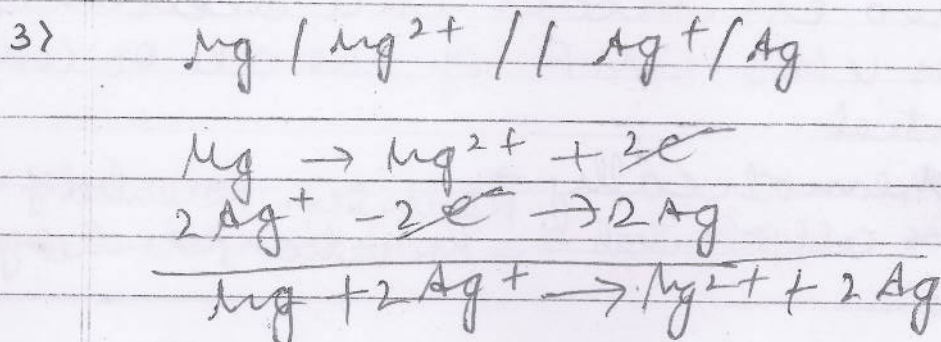
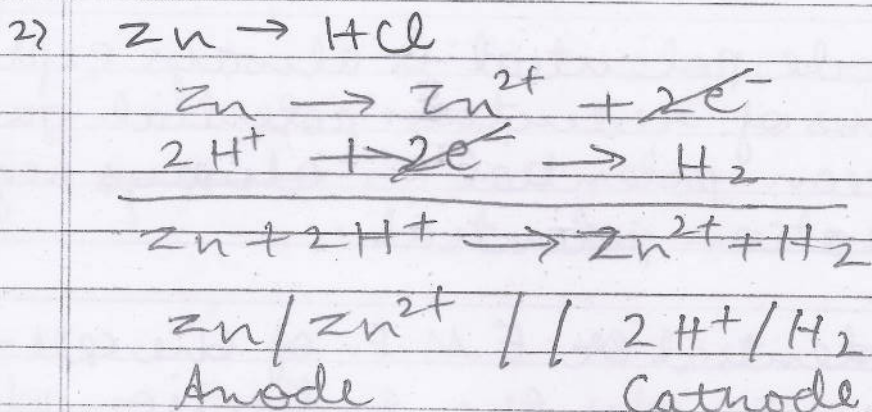
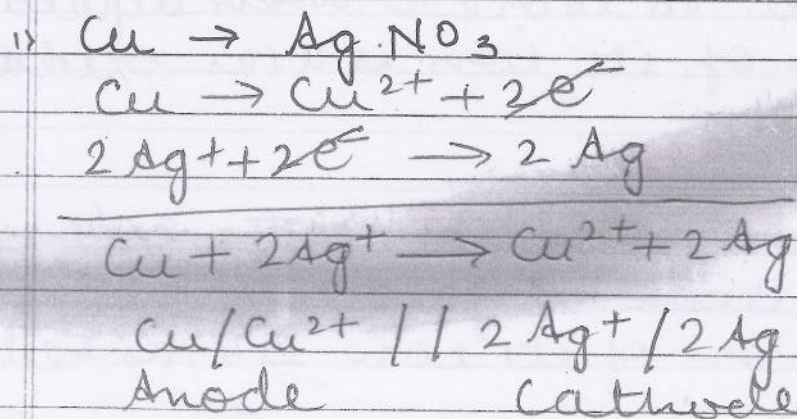
* Electrochemical cell or galvanic cell \rightarrow It is a device which convert chemical energy into electrical energy.
eg \Rightarrow Zn-CuSO₄ cell

Q. What is salt bridge and its function?
A. Salt bridge \rightarrow It is a usually an inverted U-tube filled with concentrated solution of inert electrolyte.

Functions →

- 1) Salt bridge completes the electrical circuit.
- 2) Salt bridge maintains electrical neutrality of two half cell solutions.

Representation of a electrochemical cell
 $Zn/Zn^{2+} // Cu^{2+}/Cu$
 Anode(oxidation) Cathode(reduction)



* Electrode potential \rightarrow when a metal rod or electrode is dipped in a solution of its ions a potential difference is set up between metal and the solution. This potential difference is known as electrode potential. It is represented by E .

- There are two types of electrode potential
- 1) Oxidation potential
 - 2) Reduction potential

Oxidation potential \rightarrow The tendency of an electrode to lose e^- when dipped in a solution of its ion is known as oxidation potential.

Reduction potential \rightarrow The tendency of an electrode to gain e^- when dipped in a solution of its ion is known as reduction potential.

\Rightarrow Electrode potential is always expressed in terms of reduction potential. The reduction potential is always negative of oxidation potential.

* Cell potential or E.M.F. of the cell \rightarrow The difference b/w the reduction potential of two electrodes in a electrochemical cell is known as E.M.F. of the cell or cell potential.

Mathematically \rightarrow

$$E.M.F. \text{ or cell potential} = \text{Reduction potential of Cathode} - \text{Reduction potential of anode}$$

Q. Calculate the EMF of the cell?



Given $E^{\circ}(\text{Zn}^{2+}/\text{Zn}) = -0.76 \text{ V}$

and $E^{\circ}(\text{Cu}^{2+}/\text{Cu}) = 0.34 \text{ V}$

$$\text{EMF} = 0.34 - (-0.76)$$

$$= 0.34 + 0.76$$

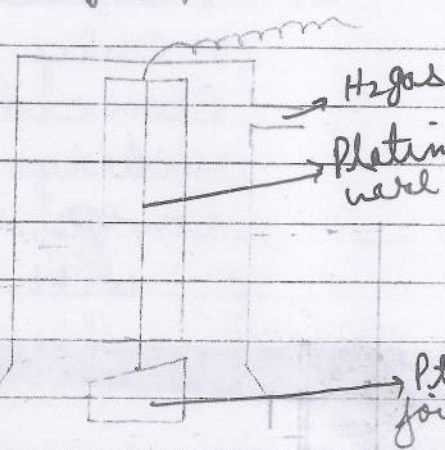
$$= 1.1$$

* Absolute value of electrode potential of an electrode cannot be determined because neither oxidation nor reduction occurs independently.

The electrode potential of an electrode is determined by connecting it with another electrode whose electrode potential is taken as standard (0). These electrodes potential of an electrode are μ/a as reference electrode.

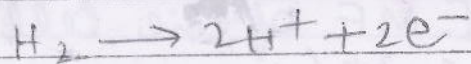
SHE (Standard hydrogen electrode) or NHE (Normal hydrogen electrode) is generally taken as reference electrode and their electrode potential is taken as zero.

* SHE or NHE \rightarrow It consists of a platinum wire having platinum foil at the bottom. This wire is sealed in a glass tube and is dipped in a solution of H^+ ions. Hydrogen gas is passed from the top to the solution of H^+ ions.

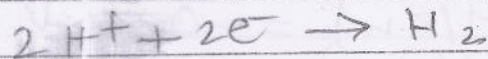


SHE OR NHE can act as a anode as well as a cathode.

when act as a anode



when act as a cathode



⇒ The electrode potential of any metal as determine w.r.t. SHE OR NHE is called standard electrode potential. It is denoted by E° .

* Electrochemical series → The arrangement of elements in a definite order of their electrode potential is called electrochemical series.

Applications of electrochemical series →

- 1) Relative strength of oxidising agent and reducing agent → The relative strength of oxidising agent and reducing agent can be determine with the help of electrode potential values in electrochemical series.
- 2) Determination of EMF of a cell → The EMF of cell can be determine with the help of electrochemical series.
- 3) Predicting the feasibility of a reaction → Only that rxn is feasible in which element having high reduction potential undergo reduction and element having lower reduction potential undergo oxidation with the help of chemical electrochemical

series we can predict the feasibility of a reaction.

* Nernst equation (Relationship b/w electrode potential or cell potential with concentration of the solution) → The relationship b/w electrode potential or cell potential with concentration of the solution can be represented by an equation known as Nernst equation.

For a general electrode rxn



then the Nernst equation can be expressed as:

$$E = E^{\circ} - \frac{RT}{nF} \ln \frac{[M(s)]}{[M^{n+}]}$$

where E = electrode potential

E° = standard electrode potential

T = Temperature

R = Gas constant

n = no. of e^{-} transfer during electrode

F = Faraday = 96500 C

\ln = natural log

$$E = E^{\circ} - \frac{2.303 RT}{nF} \log \frac{[M(s)]}{[M^{n+}]} \quad [M(s) = 1]$$

$$E = E^{\circ} - \frac{2.303 RT}{nF} \log \frac{1}{[M^{n+}]}$$

At $25^{\circ}C$ $T = 298K$

$R = 8.314 JK mol^{-1}$

$F = 96500 C$

$$E = E^{\circ} - \frac{0.059}{n} \log \frac{1}{[M^{n+}]}$$

Relationship b/w cell potential and conc. of the solution →

let us consider a cell $Zn/Zn^{2+} // Cu^{2+}/Cu$
 $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

$$E_{cell} = E(Cu^{2+}/Cu) - E(Zn^{2+}/Zn) \quad \text{--- (1)}$$

$$E(Cu^{2+}/Cu) = E^{\circ}(Cu^{2+}/Cu) - \frac{2 \cdot 303 RT}{2F} \log \frac{1}{[Cu^{2+}]}$$

$$E(Zn^{2+}/Zn) = E^{\circ}(Zn^{2+}/Zn) - \frac{2 \cdot 303 RT}{2F} \log \frac{1}{[Zn^{2+}]}$$

$$E_{cell} = E^{\circ}(Cu^{2+}/Cu) - \frac{2 \cdot 303 RT}{2F} \log \frac{1}{[Cu^{2+}]} \\ - E^{\circ}(Zn^{2+}/Zn) - \frac{2 \cdot 303 RT}{2F} \log \frac{1}{[Zn^{2+}]}$$

$$E_{cell} = E^{\circ}(Cu^{2+}/Cu) - E^{\circ}(Zn^{2+}/Zn) - \frac{2 \cdot 303 RT}{2F} \\ \left[\log \frac{1}{[Cu^{2+}]} - \log \frac{1}{[Zn^{2+}]} \right]$$

$$E_{cell} = E^{\circ}_{cell} - \frac{2 \cdot 303 RT}{2F} \log \frac{[Zn^{2+}]}{[Cu^{2+}]} \quad \left[\because \log \frac{m}{n} = \log \frac{1}{\frac{n}{m}} = \frac{n}{m} \right]$$

At $25^{\circ}C$

$$E_{cell} = E^{\circ}_{cell} - \frac{0.059}{2} \log \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

* In general for a cell $x \times n$
 $aA + bB \rightarrow cC + dD$

Nernst equation is:

$$E_{cell} = E^{\circ}_{cell} - \frac{2 \cdot 303 RT}{nF} \log \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

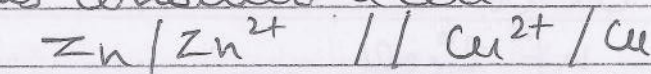
At $25^{\circ}C$ or $298K$

Nernst equation is:

$$E_{cell} = E^{\circ}_{cell} - \frac{0.059}{n} \log \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Relationship b/w E.M.F. of a cell and equilibrium constant \rightarrow

Let us consider a cell



$$K_c = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} \quad [\because [\text{Zn}] = 1 \text{ and } [\text{Cu}] = 1]$$

Now the Nernst equation for above cell

is n :

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{2.303RT}{2F} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

At equilibrium E.M.F. = 0

$$0 = E^{\circ}_{\text{cell}} - \frac{2.303RT}{2F} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

$$E^{\circ}_{\text{cell}} = \frac{2.303RT}{2F} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

$$\text{or } E^{\circ}_{\text{cell}} = \frac{2.303RT}{2F} \log K_c$$

$$\text{where } K_c = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} \text{ at } 298 \text{ K}$$

$$E^{\circ}_{\text{cell}} = \frac{0.059}{n} \log K_c \text{ at } 298 \text{ K}$$

$$\text{or } \log K_c = \frac{n E^{\circ}_{\text{cell}}}{0.059} \text{ at } 298 \text{ K}$$

in general

$$E^{\circ}_{\text{cell}} = \frac{2.303RT}{nF} \log K_c$$

Relationship b/w Gibbs free energy change and E.M.F. of the cell \rightarrow

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

ΔG = Gibbs free energy change

eg \Rightarrow Calculate ΔG change (max. work) for a cell

$$\text{Cu} / \text{Cu}^{2+} // \text{Ag}^{2+} / \text{Ag}$$

Given $E^\circ(\text{Ag}^+ / \text{Ag}) = 0.80 \text{ V}$
 $E^\circ(\text{Cu}^{2+} / \text{Cu}) = 0.34 \text{ V}$
 $E^\circ_{\text{cell}} = 0.80 - 0.34$
 $= 0.46$

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

$$\Delta G = -2 \times 96500 \times 0.46$$

$$\Delta G = -193000 \times 0.46$$

$$\Delta G = -88780.00$$

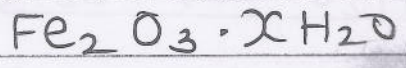
* Difference b/w electrochemical cell and electrolytic cell \rightarrow

	Electrochemical cell	Electrolytic cell
1)	It is a device which converts chemical energy into electrical energy.	It is a device which converts electrical energy into chemical energy.
2)	In this case, the redox reaction is spontaneous and is responsible for the production of electrical energy.	The redox rxn is non-spontaneous and electrical energy is supplied to make the rxn to occur (i.e. chemical change to take place)
3)	The two half cells are set up in different containers and are connected through salt bridge.	Both the electrodes are placed in the solution electrolyte in the same container.
4)	Anode is $-ve$ and Cathode is $+ve$ electrode	Anode is $+ve$ and Cathode is $-ve$ electrode.
5)	The e^- move from anode to \ominus Cathode.	The e^- are supplied by external battery and enter through Cathode. & leave from anode.

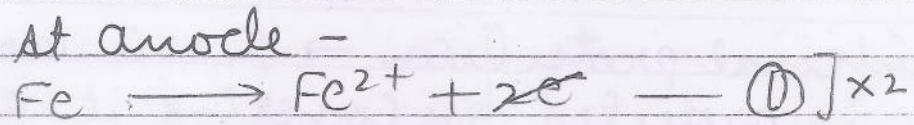
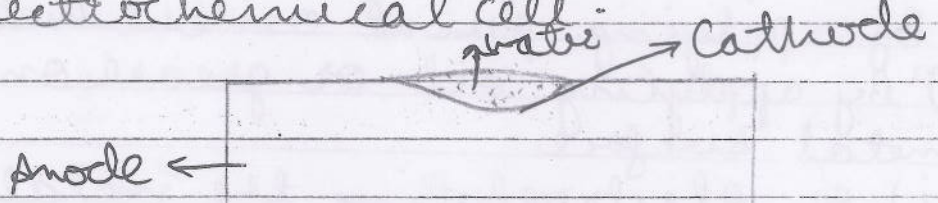
* **Corrosion** → A process of deterioration of metal due to its rxn with air and water present in the environment is called as corrosion.

If metal is iron then corrosion is called as rusting.

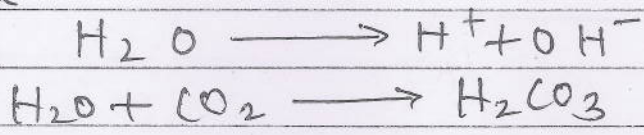
Rusting is a process of deterioration of iron due to its rxn with air and water present in the environment. Chemically rust is hydrated ferric oxide -

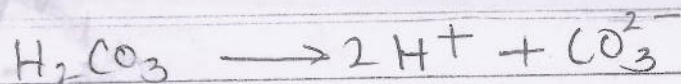


Mechanism of rusting of iron → Rusting of iron can be explained by an electrochemical theory. According to this theory pure iron metal behaves as an anode. The impurity or uneven surface behaves as a cathode and the whole arrangement behaves as an electrochemical cell.

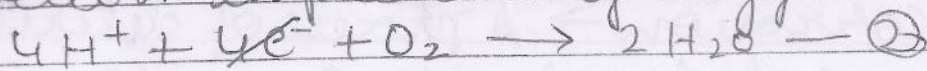


At cathode, there is H^+ ions due to following rxn





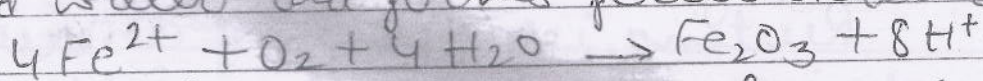
The H^+ ions present at Cathode undergo reduction in presence of oxygen



\therefore Overall rxn



The Fe^{2+} ion formed in above equation undergoes oxidation in presence of oxygen and water and forms ferric oxide (Fe_2O_3)



Ferric oxide react with water molecules and forms hydrated ferric oxide i.e. rust as a product.



Prevention of corrosion \rightarrow

Corrosion can be prevented by following methods

- 1) **Barrier protection** \rightarrow This is the simplest method for prevention of corrosion. In this method a barrier is placed b/w the atmosphere and metal. This barrier can be placed by following ways \rightarrow
 - i) By applying paint on the metal surface
 - ii) By applying oil or grease on the metal surface.
 - iii) By electroplating the metal surface.
- 2) **Sacrificial protection** \rightarrow In this method iron is protected by covering it with the layer of some other metal which is more active than iron.

Zinc is generally used to cover the iron surface. This process of covering of the iron surface with Zn metal in order to prevent it from rusting is called galvanisation.

3) Electrical protection → This method is used for protecting iron articles which are in contact with water such as underground pipes. In this method the iron article is connected with some other metal which has lower reduction potential than iron.

4) Using anti-rust solutions → Corrosion or rust can be prevented by applying anti rust solutions such as alkyl phosphate and alkyl chromates.

* Some commercial cells / batteries →
Battery → It is an arrangement of electrochemical cells that can be used as a source of direct electric current at a constant voltage.

A useful battery should also fulfil the following requirements →

- i) It should be light and compact so that it can be easily transported.
- ii) It should have reasonably long life both when it is being used and when it is not used.
- iii) The voltage of the battery should not vary.

appreciably during its use.

Types of batteries → There are mainly two types of batteries:

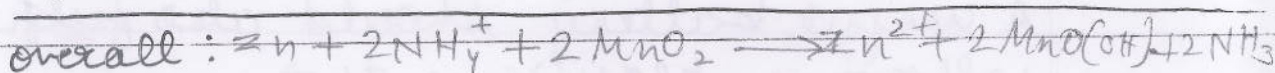
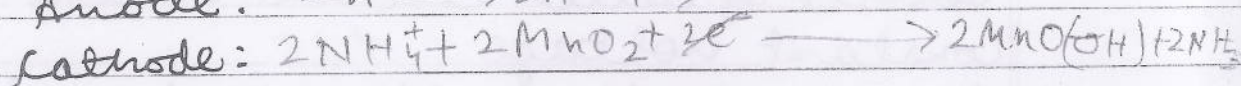
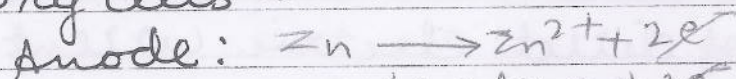
- 1) Primary batteries or cells
- 2) Secondary batteries or cells

Primary batteries or cells → In these cells, the electrode reaction cannot be reversed by an external electric energy source. In these cells reaction occurs only once and after use they become dead. Therefore they are not chargeable.
eg. → dry cell, mercury cell.

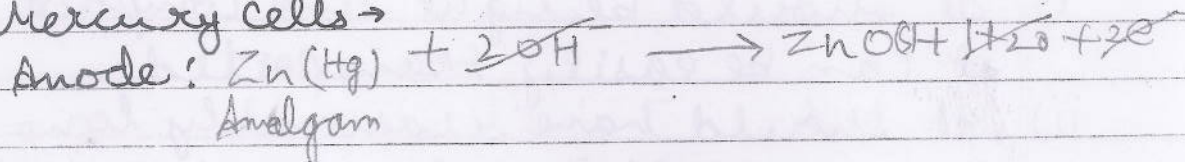
Secondary batteries or cells → In these cells, the rxn can be reversed by an external electric energy source. Therefore, these cells can be recharged by passing electric current and used again and again. These are also called storage cells.
eg. → lead storage battery and nickel-cadmium storage cell.

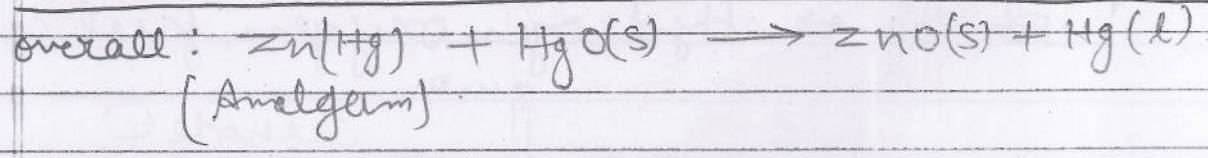
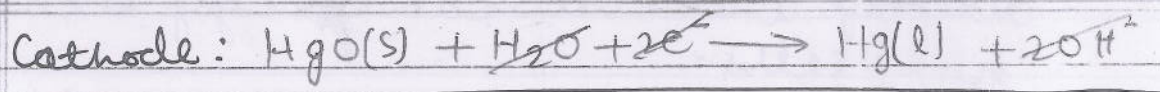
eg. of primary cells →

1) Dry cells →



2) Mercury cells →

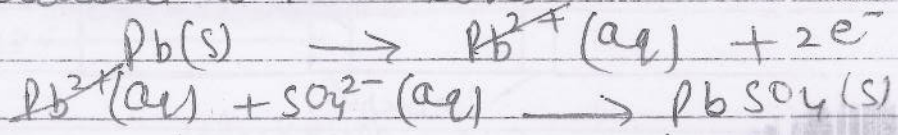




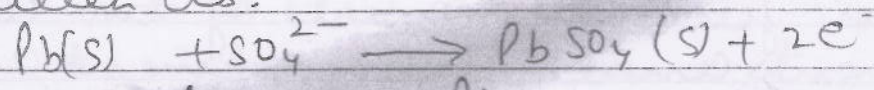
eg. of secondary cells →

1) Lead storage cell

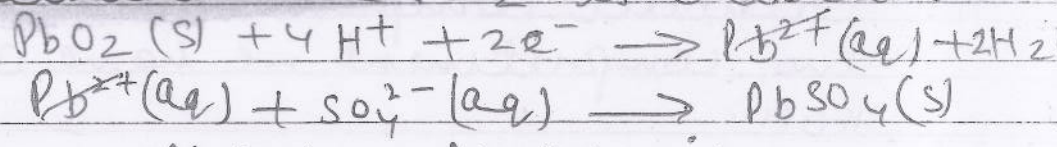
At anode: The lead loses two electrons and is oxidised to Pb^{2+} ions



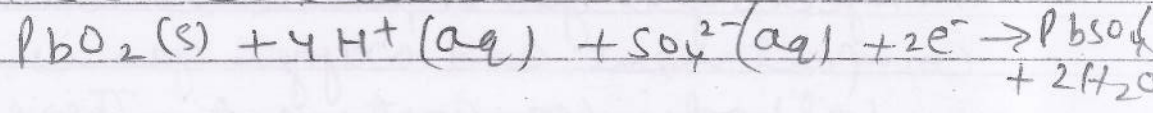
The overall ~~rxn~~ anode rxn may be written as:



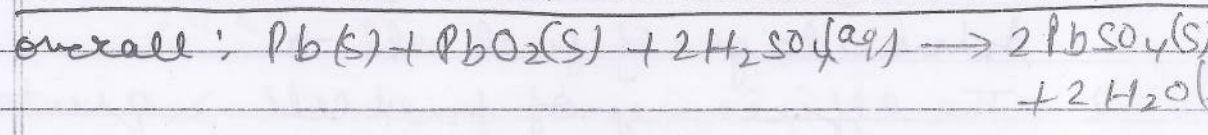
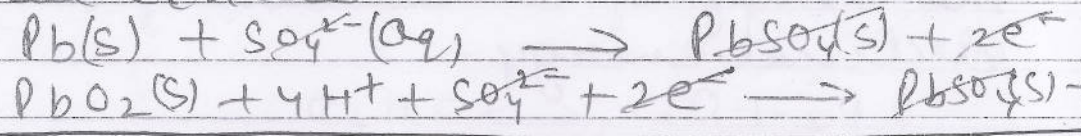
At cathode: The PbO_2 is reduced as:



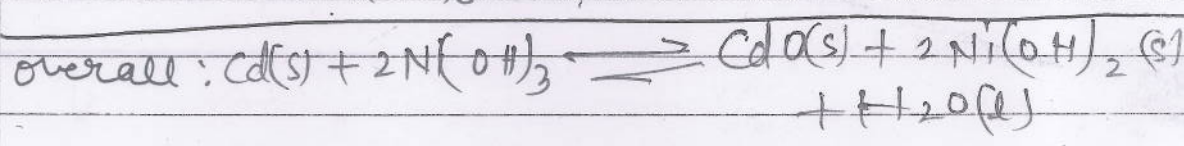
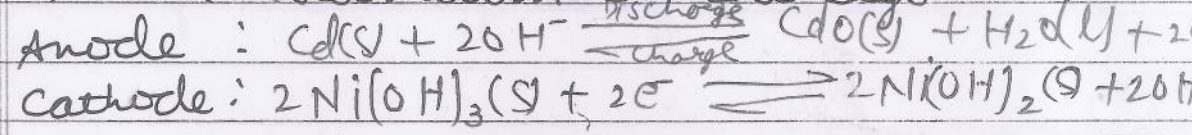
The overall cathode rxn is



Thus the complete electrode rxn and overall cell rxn are:



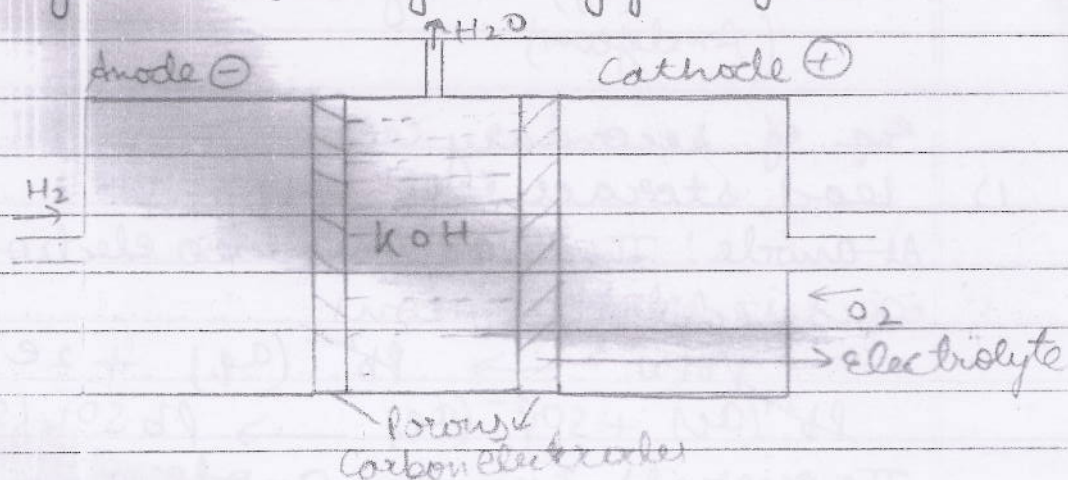
2. Nickel cadmium ~~cell~~ storage cell →



NH_3
-
 NH_3

* Fuel cells \Rightarrow The cells which convert the energy of combustion of fuel (Hydrogen, methane) directly into electrical energy are called fuel cells.

For eg. \Rightarrow Hydrogen oxygen fuel cell



This fuel cell consists of two compartments anodic compartment and cathodic compartment b/w these two compartment we have a solution of KOH.

Hydrogen gas is passed in anodic compartment and oxygen gas is passed in cathodic compartment. These gases react with each other and convert the energy of combustion of fuel into electrical energy.

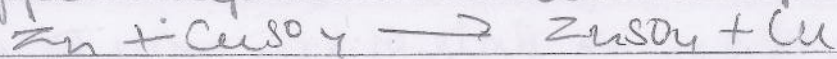


Advantages of fuel cell \rightarrow

- 1) The efficiency of fuel cell is quite high.
- 2) It gives continuous energy.
- 3) Its working is pollution free.

Q. Can we store Cu^{sulphate} solution in a Zn pot?

Ans. No, because zinc is more reactive than copper and it will displace copper from copper sulphate solution.



As a result zinc will gradually pass into the sol. and Cu will be precipitated.

Q. Can zinc spoon be used to stir CuSO₄ sol.

Ans. No, we can't use Zn spoon to stir CuSO₄ sol. because Zn spoon reacts with CuSO₄ solution.



Q. Can Cu spoon be used to stir ZnSO₄ sol.

Ans. Yes, we can use Cu spoon to stir ZnSO₄ solution because Cu does not react with ZnSO₄ solution.

Q. What is cell constant and its unit?

Ans. It is the length per unit area i.e. $\frac{l}{a}$ units $\rightarrow \text{cm}^{-1}$ or m^{-1}