

**UNIT-III**

**ELECTRO**

**CHEMISTRY**

**,BATTERIES**

**AND CORROSION**

## Introduction & scope:

Electrochemistry is the study of the processes involved in the interconversion of electrical energy & chemical energy. It concerned with the electrical aspects of chemical reactions particularly electrolysis & cells.

### Concept of electrochemistry

In brief it deals with the chemical applications of electricity. Broadly we can classify the cells as electrolytic cells & galvanic cells. In electrolytic cells the conversion of electrical energy to chemical energy takes place. whereas the chemical energy is converted to electrical energy in a galvanic cell.

### conductance - Electrolyte in solution

#### Electric Conductors

Electric current is the flow of electrons generated by a battery when the circuit is completed. Any substance which allows the passage of electric current through it is called electric conductor.

e.g.: All metals, graphite, aq solution of acids & bases

Electric conductors are 2 types. They are

- 1) Metallic conductors
- 2) Electrolytic conductors.

## Metallic conductors

Metallic conductors are which conducts the electricity by free, mobile, valency electrons. They involve flow of electrons & there is no chemical decomposition during conductance.

Eg: Metals, alloys, certain solid states etc.

## Electrolytic conductors:

These are aq. solutions (or) fused electrolytes which liberate ions & conduct electricity involving the movement of ions, resulting in its chemical decomposition.

Eg: Acids & Bases.

## Electrolysis:

The process of decomposition of an electrolyte by passing electric current through its solution is called as electrolysis, which is carried in an apparatus called the electrolytic cell.

The cell contains aq. solution of an electrolyte in which two metallic rods (or) electrodes are dipped which are connected to a battery.

## Faraday's laws of electrolysis:

It tells that the relationship b/w quantity of electrical charge passing through an electrolyte & the amount of the substance deposited at the electrode. There are 2 Faraday's law of electrolysis.

### Faraday 1<sup>st</sup> law:

According to the Faraday's 1<sup>st</sup> law of electrolysis • The amount of the substance ( $N$ ) which is

deposited at the electrode is directly proportional to the quantity of the electrical charge ( $Q$ ).

$$\text{i.e. } \boxed{H \propto Q}$$

Here  $w \rightarrow$  Amount of the substance

$Q \rightarrow$  Quantity of the charge.

$$H = zQ$$

[But  $Q = it$ ]

$$\boxed{H = zit}$$

Here  $z \rightarrow$  Electrochemical equivalent constant.

When 1 amp of current flows through the conductor per sec the quantity of electricity will be 1 coulombs.

$$\text{i.e. } i=1 \text{ & } t=1 \text{ sec} \Rightarrow \boxed{H=z}$$

It has been found that the quantity of electricity required to liberate an 1 gm of equivalent of substance is 96,500 C, this quantity of electrical charge is known as Faraday

$$\therefore \boxed{1F = 96,500 C}$$

### Faraday II - law of electrolysis

~~Rule~~

According to the Faraday II-law of electrolysis when the same quantity of electrical charge is passing through the different electrolytes, then the amount of different substance deposited on the electrodes is  $[H]$  directly proportional to their equivalent weight (E)

$H_1$  &  $H_2$  are the amounts of different substances

$E_1$  &  $E_2$  are their equivalent weights.

$$\therefore H_1 \propto E_1 \text{ & } H_2 \propto E_2$$

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

But according to the Faraday's 1<sup>st</sup> law of electrolysis

$$W_1 = Z_1 i t$$

$$W_2 = Z_2 i t$$

$$\therefore \frac{E_1}{E_2} = \frac{Z_1 i t}{Z_2 i t}$$

$$\boxed{\frac{E_1}{E_2} = \frac{Z_1}{Z_2}}$$

i.e  $\boxed{E \propto z}$

∴ Electrochemical equivalent constant is directly proportional to their equivalent weights.

$$\boxed{E = F \times z}$$

$$\boxed{\text{Here } 1F = 96,500 C}$$

$$\therefore \boxed{E = 96,500 C \times z}$$

=

SUB:

NAME OF THE FACULTY:

TOPIC:

Specific Resistance ( $\rho$ ):

The Resistance of the conductor ( $R$ ) is directly proportional to the length of the conductor ( $l$ ) and inversely proportional to the area of cross section ( $a$ )

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

$$\text{i.e } R = \rho \cdot \frac{l}{a}$$

Here  $\rho$  is called Specific Resistance.

$$\therefore \boxed{\rho = \frac{R \cdot a}{l}}$$

$$\text{units: } \rho = \frac{\text{ohm} \cdot \text{cm}^2 / \text{m}}{\text{cm/m}} \Rightarrow \boxed{\rho = \text{ohm} \cdot \text{cm/m}}$$

Specific Conductance ( $K$ )

Reciprocal of the Specific Resistance is called Specific Conductance ( $K$ ).

$$\therefore K = \frac{1}{\rho} \quad [\text{But } \rho = \frac{R \cdot A}{l}]$$

$$\therefore K = \frac{1}{\frac{R \cdot A}{l}}$$

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

*K.C.K*

Here  $\frac{l}{A}$  is called cell constant & it is represented by ' $x$ '  $\therefore \left\{ x = \frac{l}{A} \right\}$ . units  $\Rightarrow x = \frac{\text{cm}}{\text{cm}^2} = \text{cm}^{-1}$

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

## Equivalent Conductance ( $\lambda_{eq}$ )

Equivalent conductance which is defined as the conductance of the solution which containing the 1 gram equivalent wt of the electrolyte is called Equivalent Conductance ( $\lambda_{eq}$ )

$$\therefore \lambda_{eq} = \frac{1000}{c} \times K$$

Here  $K$  is specific conductance.

$$\therefore [K = \frac{1}{R} \times \frac{1}{A}]$$

$$\therefore \lambda_{eq} = \frac{1000}{c} \times \frac{1}{R} \times \frac{1}{A}$$

units:  $\lambda_{eq} = \frac{1}{R} \times \frac{1}{A} \times V \times \frac{1}{c}$   
 $= \frac{1}{\text{Ohm}} \times \frac{\text{cm}}{\text{cm}^2} \times \text{cm}^2 \times \frac{1}{\text{eq}}$

$$\lambda_{eq} = \text{Ohm}^{-1} \text{cm}^2 \text{eq}^{-1}$$

## Molar conductance ( $\lambda_m$ )

Molar conductance which is defined as the conductance of the solution which containing the 1 mole of electrolyte called molar conductance ( $\lambda_m$ )

$$\lambda_m = \frac{100c}{c} \times K$$

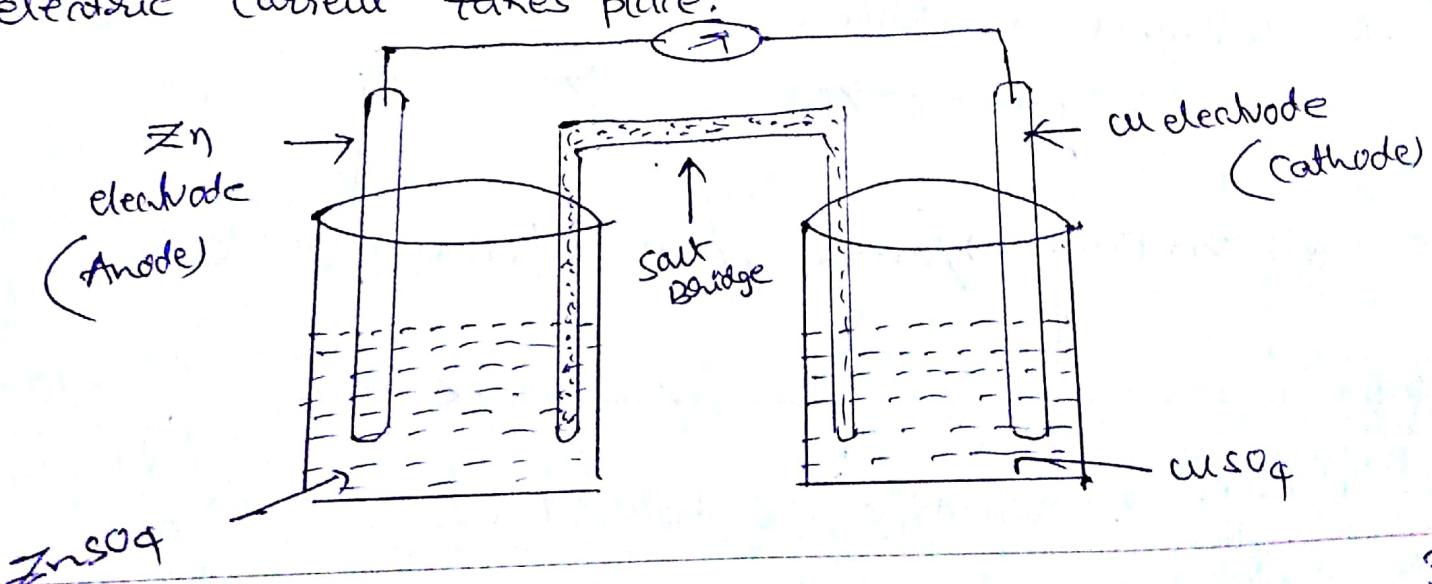
units:  $\lambda_m = \text{Ohm}^{-1} \text{cm}^2 \text{mol}^{-1}$

$\equiv$

## "Electrochemical cells": EMF

An electrochemical cell is a device for converting chemical energy into electrical energy. A Redox reaction is utilised to get electrical energy. An electrochemical cell is also commonly referred to as voltaic (or) galvanic cell. The electromotive force (emf) of such a cell is directly proportional to the intensity of the chemical reaction taking place in it.

The practical application of an electrochemical cell (or galvanic cell) is Daniell cell. It consisting of Zn electrode dipping in  $ZnSO_4$  solution (where oxidation reaction takes place) & A copper rod dipping in  $CuSO_4$  solution (where reduction takes place). In other words, each electrode may be regarded as a "half cell". The two solutions are separated by a salt bridge (or) semi permeable membrane & these prevent the diffusion of the two liquids but allows the passage of ions through it, when the flow of electric current takes place.



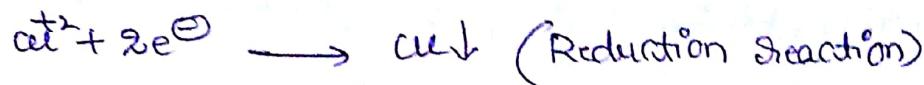
When the two metal electrodes are connected the flow of electric current takes place. Actually what happens is that Zn passes into the solution as  $Zn^{+2}$  ion liberating two electrons because of its higher oxidation potential than copper.

At anode:

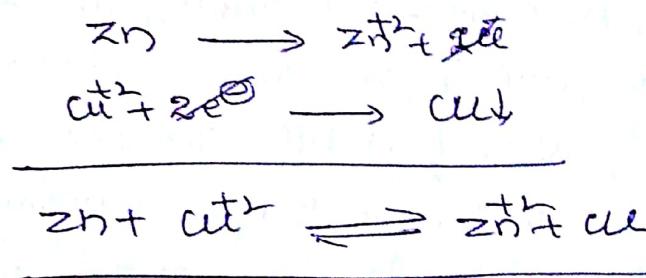


The electrons thus liberated travel along the external circuit to the copper electrode where copper ions ( $Cu^{+2}$ ) gain these electrons & converted to metallic copper.

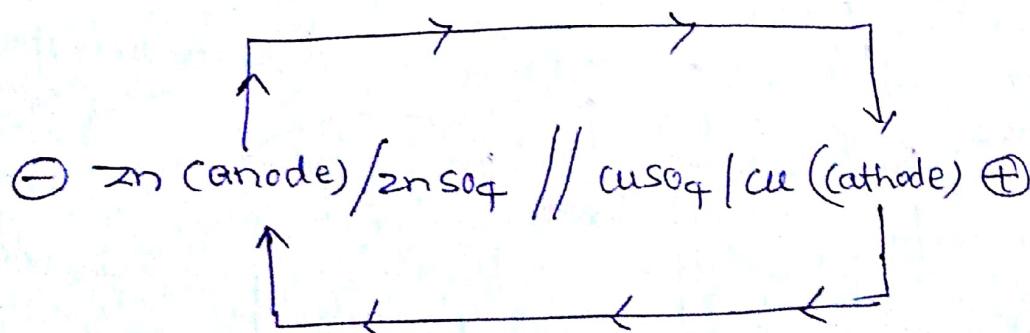
At cathode:



The movement of the electrons from Zn to Cu produces a current in the circuit & the net chemical reaction can be represented as follows.



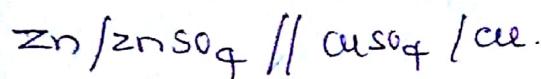
∴ The total cell along with the flow of electrons is described as follows.



Direction of flow of e's.

### Cell notation:

The total cell according to the convention is represented as



The emf of the Daniell cell is 1.09 V (or) 1.1V

∴ The emf of the galvanic cell is  $E_{\text{cell}} = E_R - E_L$

### Single electrode potential

Each electrochemical cell is made up of two electrodes, at one electrode electrons are evolved and the other are used up. Each electrode which is dipped in its salt solution is called "Half cell".

The potential of half cell, i.e. the potential difference b/w the metal and its salt solution in which it is dipped is called single electrode potential.

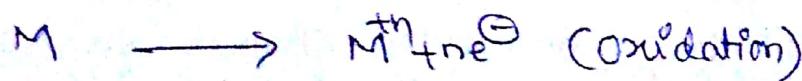
### Standard Electrode potential

It is the measure of tendency of a metallic electrode to lose (or) gain electrons, when it is in contact with a solution of its own salt of 1M concentration at 25°C.

### Electrode potential

A metal (M) consisting of Metal ions ( $M^{+n}$ ) with valence electrons. When the metal (M) is placed in a solution of its own salt, any one of the following reactions will occurs

→ A positive metal ions may pass into the solution



→ positive metal ions from the solution may deposit over the metal.



### Types of Electrodes:

The tendency of a metal to lose  $e^-$  or gain the  $\ominus$  are called as Electrode.

They are  
Electrodes are classified into 2 types

- 1) Reference electrodes
- 2) Ion-selective electrodes.

### Reference electrodes:

The determination of the absolute value of the potential of an electrode is not possible since every circuit necessarily contains two electrodes. This difficulties has however been overcome by using an electrode of known potential as reference electrode. In practice the electrode whose potential is to be determined is combined with the reference electrode.

By the help of a reference electrode we can determine the emf value of the other half cell.

- 1) Standard Calomel electrode (SCE)
- 2) Quinhydrone electrode
- 3) Ion-selective electrode or Glass electrode.

### \* Calomel Electrode (or) Mercury, Mercurous Chloride Electrode

Calomel electrode is very widely used and it is the mercury-mercurous chloride electrode.

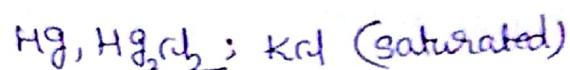
The potential of the calomel electrode has been determined by connecting it to standard hydrogen electrode. The potential value has been found to vary with the concentration of the KCl solution used in calomel electrode.

for 0.1 (Dei normal) N KCl — 0.3335 V (potential)

1 N KCl — 0.2810 V

Saturated KCl — 0.2428 V

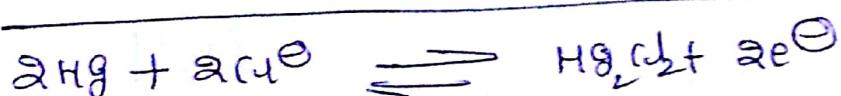
Calomel electrode represented as



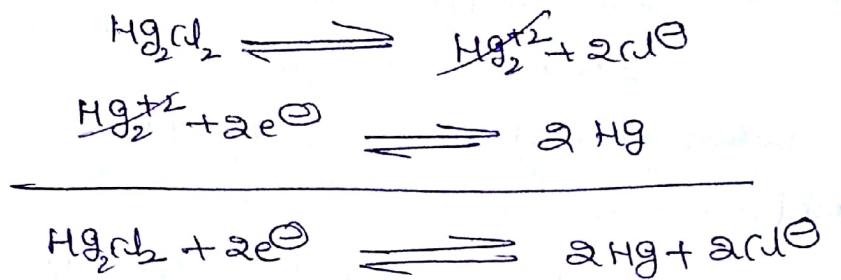
### Structure:

The calomel electrode consisting of wide tube with narrow side tube. A pt core is dipped into liquid mercury which is covered with solid  $\text{Hg}_2\text{Cl}_2$ . The tube is filled with saturated KCl solution and the side tube containing KCl solution act as a salt bridge which connects this electrode to any other electrode. Here pt core used for making electrical current, in electrolytic solution.

If this electrode acts as anode i.e when oxidation reaction takes place at the electrode. The liquid metal (Hg) gives mercurous ( $\text{Hg}_2^{+2}$ ) ions into solution with the liberation of electrons and these  $\text{Hg}_2^{+2}$  ions are combined with the  $\text{Cl}^-$  ions from KCl solution to form  $\text{Hg}_2\text{Cl}_2$ .



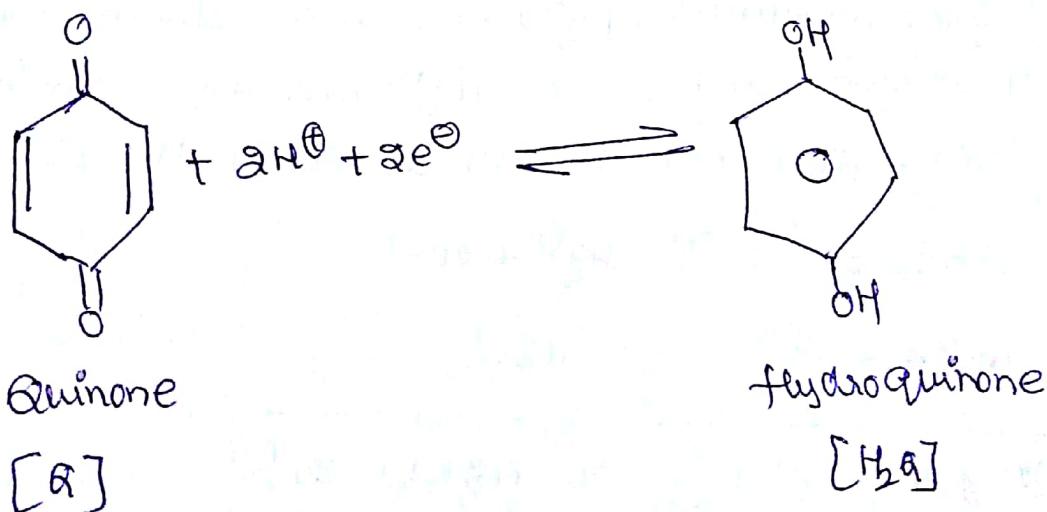
On the other hand when the electrode acts as cathode reduction reactions takes place. The  $\text{Hg}_2^{+2}$  ions are given by the  $\text{Hg}_2\text{Cl}_2$  would be discharged at the electrode. Hence more & more of  $\text{Hg}_2\text{Cl}_2$  pass into solution which results in an increase in the concentration of  $\text{Cl}^-$  ions.



### "Quinhydrone Electrode"

This type of electrode involves redox reaction b/w Quinone & hydroquinone, when hydroquinone & quinone are added to a sample solution. A pH dependent oxidation - Reduction reactions are formed in b/w quinone & hydroquinone.

This redox potential is measured with a pt electrode. When pt electrode is immersed in a solution containing quinhydrone & quinone, a potential is developed.



→ potential of anhydride electrode is

$$E_Q = E_Q^{\circ} - \frac{2.303RT}{2F} \log \frac{[H_2A]}{[A][H^+]^2}$$

and if quinone and hydroquinone are taken in equimolar concentration, then  $[A] = [H_2A]$

$$E_Q = E_Q^{\circ} - \frac{2.303RT}{2F} \log \frac{1}{[H^+]^2}$$

$$E_Q = E_Q^{\circ} - \frac{2.303RT}{2F} \log (H^+)^{-2}$$

$$E_Q = E_Q^{\circ} - \frac{2.303RT}{2F} (-2) \log (H^+)$$

$$E_Q = E_Q^{\circ} - \frac{2.303RT}{F} (-\log (H^+))$$

$$\boxed{E_Q = E_Q^{\circ} - 0.05916V \times pH}$$

→ An examination of this equation shows that the above system behave as a form of hydrogen electrode. Hence this electrode can be employed for measuring pH of solutions as an indicator electrode.

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High  
Note

## Glass Electrode (or) Ion-Selective electrode

This electrode possess the ability to respond only certain specific ions, thereby developing a potential with respect to that species only in a mixture & ignoring the other ions totally.

The potential developed by an ion-selective electrode depends up on the concentrations of Species.

### Glass electrode:

Glass electrode selective for only  $\text{H}^+$  ions for a particular type of glass, the potential difference varies with the  $\text{H}^+$  ion concentration & it is given by the expression.

i.e

$$E_G = E_G^\circ - \frac{RT}{nF} \ln(\text{H}^+)$$

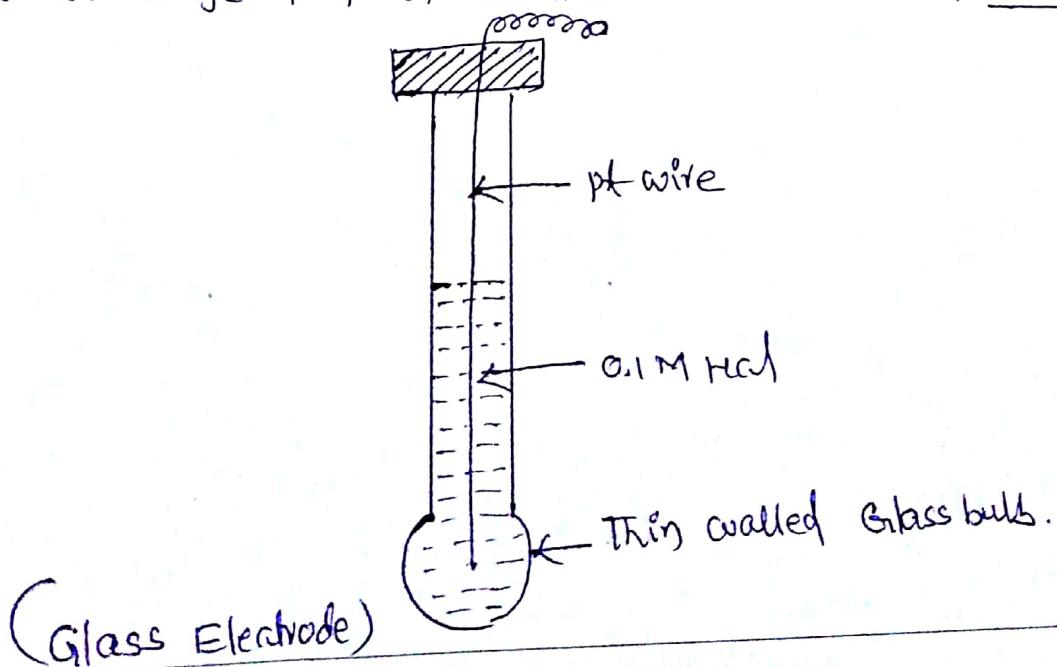
$$E_G = E_G^\circ - \frac{RT}{F} \ln(\text{H}^+) \quad [\because n=1; \text{H}^+]$$

$$E_G = E_G^\circ - 0.05916 \log(\text{H}^+)$$

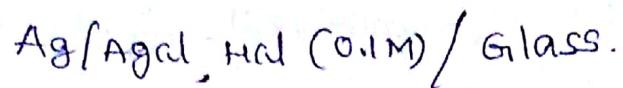
$$E_G = E_G^\circ + [0.05916 (-\log(\text{H}^+))]$$

$$\boxed{E_G = E_G^\circ + 0.05916 \times \text{pH}} \quad \text{at } 25^\circ\text{C}.$$

over a range of pH of the test-solution from 0 to 10



→ A glass electrode consisting of thin-walled glass bulb containing AgCl-coated Ag electrode (or) Simply a pt electrode in 0.1M HCl. The glass electrode may be shown as.

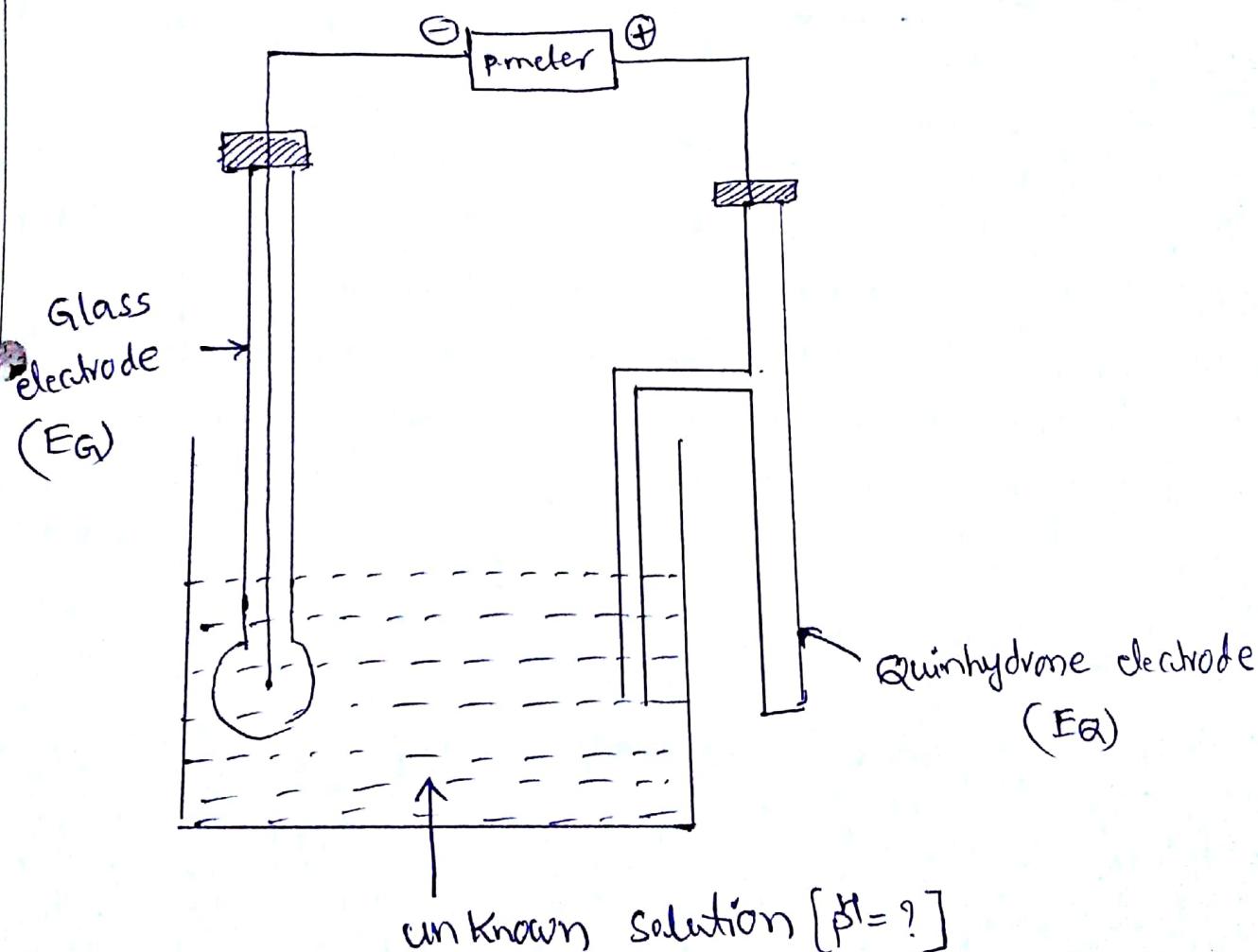


HCl in the bulb furnishes a constant H<sup>+</sup> ion concentration

### Application:

Determination of pH of unknown solution by using Glass electrode

In order to determine the pH of a solution, the glass electrode is placed in the solution under-test & this half cell is coupled with quinhydrone electrode.



$$E_{\text{cell}} = E_R - E_L$$

$$E_{\text{cell}} = E_Q - E_G$$

$$E_{\text{cell}} = E_Q^0 - 0.05916 \times p^H - E_G^0 - 0.05916 \times p^H$$

$$E_{\text{cell}} = E_Q^0 - E_G^0 - 0.05916 (p^H + p^H)$$

$$E_{\text{cell}} = E_Q^0 - E_G^0 - 0.05916 \cdot 2p^H$$

$$0.05916 \cdot 2p^H = E_Q^0 - E_G^0 - E_{\text{cell}}$$

$$2p^H = \frac{E_Q^0 - E_G^0 - E_{\text{cell}}}{0.05916}$$

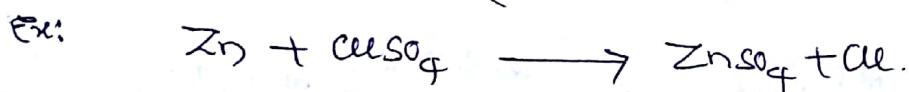
$$\boxed{p^H = \frac{E_Q^0 - E_G^0 - E_{\text{cell}}}{0.05916 \times 2}}$$

## Electrochemical series and its applications.

When the metals are arranged in the order of increasing reduction potentials (or) decreasing oxidation potentials which are then determined with respect to one-molar solutions of their ions and measured on the hydrogen scale, along series (or) list, resulted is called electrochemical series.

The higher a metal is in the series, the greater is its tendency to be oxidised. The metals high up in the series are strong reducing agents and their ions are stable whereas those near the bottom of the series are inactive, stable metals, and their ions are easily reduced to metals. Any metal can reduce any cation below it in the series from its one molar solutions.

The metals above the hydrogen in the series can easily displace the metal below the hydrogen in the series from its salt solution, because  $E^\circ$  of  $Zn^{+2}/Zn$  (-0.763V) is less than  $E^\circ$  of  $Cu^{+2}/Cu$  (0.337V).



∴ The metals above the hydrogen in the series can be easily oxidised hence they undergo corrosion easily.

Standard oxidation potentials at 25°C.

Electrode	Half cell reaction	$E^\circ$ volts.
Li/Li <sup>+</sup>	$\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$	+ 3.045 V
K/K <sup>+</sup>	$\text{K} \rightarrow \text{K}^+ + \text{e}^-$	+ 2.925 V
Ca/Ca <sup>2+</sup>	$\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$	+ 2.87
Na/Na <sup>+</sup>	$\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$	+ 2.714
Mg/Mg <sup>2+</sup>	$\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$	+ 2.37
Zn/Zn <sup>2+</sup>	$\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$	+ 0.763
Pt/H <sub>2</sub> /H <sup>+</sup>	$\text{H}_2 \rightarrow 2\text{H}^+ + 2\text{e}^-$	0.0000 V
Ce/Ce <sup>4+</sup>	$\text{Ce} \rightarrow \text{Ce}^{4+} + 2\text{e}^-$	- 0.337 V
Ag/Ag <sup>+</sup>	$\text{Ag} \rightarrow \text{Ag}^+ + \text{e}^-$	- 0.799 V
Au/Au <sup>3+</sup>	$\text{Au} \rightarrow \text{Au}^{3+} + 3\text{e}^-$	- 1.50 V

### Applications:

The electrochemical series give more information on

1. The relative corrosion tendencies of the metals & alloys.
2. Relative ease of oxidation (or) reduction of metals.
3. Replacement tendency of metals.
4. Calculating the equilibrium constant as given below.

$$E^\circ = \frac{RT}{nF} \ln K_{\text{eq}} = \frac{2.303 RT}{nF} \log_{10} K_{\text{eq}}$$

$$\log_{10} K_{\text{eq}} = \frac{nFE^\circ}{2.303 RT}$$

$$= \frac{nFE^\circ}{0.0591} \text{ at } 25^\circ\text{C}$$

## Potentiometric titration:

Potentiometric titration is very important application of emf measurement. i.e. the potential of an electrode is mainly depends upon the concentrations of the ions.

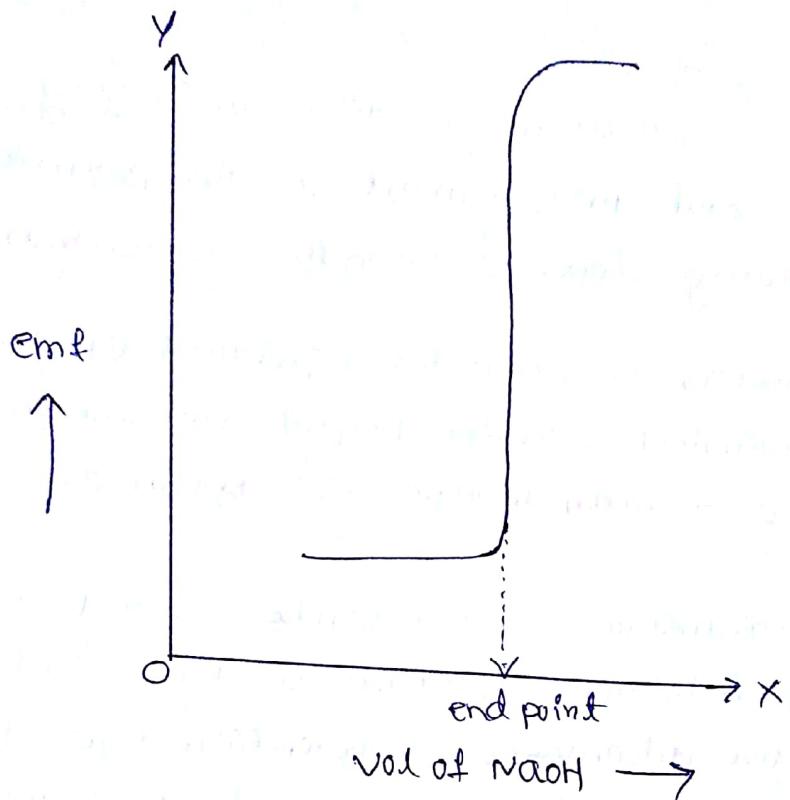
A titration in which the equivalent or end point of a reaction is determined with the help of measurement of the potentials of the reaction mixture is known as potentiometric titration.

Potentiometric titration can be carried out in coloured solutions while indicators can't be used in such cases. This is one of the advantages of potentiometric titrations. But ordinary titrations carried out by the help of indicators.

### (i) Acid-Base Titration:

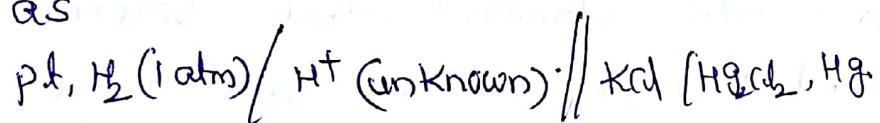
The acid solution whose strength has to be determined is taken in a beaker and the hydrogen electrode & Calomel electrode were dipped in the solution. The electrodes were connected to the potentiometer and the emf measured. A known volume, standard alkali solution is added from a burette, stirred thoroughly and the emf of the cell is recorded. Like this 10-15 readings are recorded, by repeating the procedure of the addition of standard alkali. The volume of alkali added is plotted against emf observed as shown in the diagram (Graph).

The steepest portion of the curve indicates the equivalent point of the titration. The change of potential will be slow at first, but at equivalent point the point change will be sharp.



### Determination of pH by emf method.

The emf of a solution depends upon the concentration of H<sup>+</sup> ions (or) pH of the solution. A hydrogen electrode containing solution of unknown pH is paired with a standard calomel electrode. The complete cell may be represented as



The emf of the above cell is measured by potentiometer and the pH of the unknown solution can be calculated as follows.

Total emf of the cell is

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

$$E_{\text{cell}} = E_{\text{cat}} - E_{\text{pt electrode}} (\text{f}_{\text{B electrode}})$$

$$E_{\text{cell}} = 0.2422 - (-0.05916 \times \text{pH})$$

$$E_{\text{cell}} = 0.2422 + 0.05916 \times \text{pH}$$

$$\boxed{\text{pH} = \frac{E_{\text{cell}} - 0.2422}{0.05916 \text{ V.}}}$$

# BATTERIES

## Introduction & scope:

Battery is an electrochemical cell which converts chemical energy into electrical energy called battery

(Or)

When two (or) more electrochemical cells are electrically interconnected each of which containing two electrodes & electrolyte is called as battery.

## Classification

Batteries are broadly classified into 2 types depending on their recharging capabilities. They are

- 1) primary batteries
- 2) Secondary batteries.

### primary Batteries: [Non-Rechargeable Batteries]

The primary batteries in which the cell reaction is not reversible i.e when the reactants have been converted into products, no electricity is produced & the cell becomes dead & can't be used after that. These batteries are used as a source of D.C power.

Ex: Dry cells, Mercury cell.

### Secondary Batteries: [Rechargeable Batteries]

The batteries in which the cell reaction is reversed by passing direct current through a in opposite direction these can be used through a large number of cycles of discharging & charging.

- Eg: 1) Lead-acid storage battery
- 2) Ni-cd battery
- 3) Li-cells.

## Lithium Battery: (or) Lithium cell

Lithium cell is a primary cell, lithium metal will acts as anode here

it has the following three components.

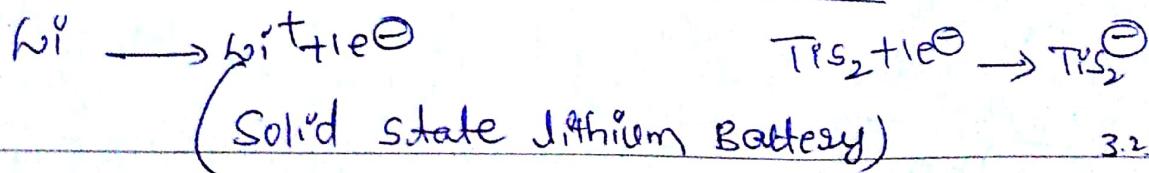
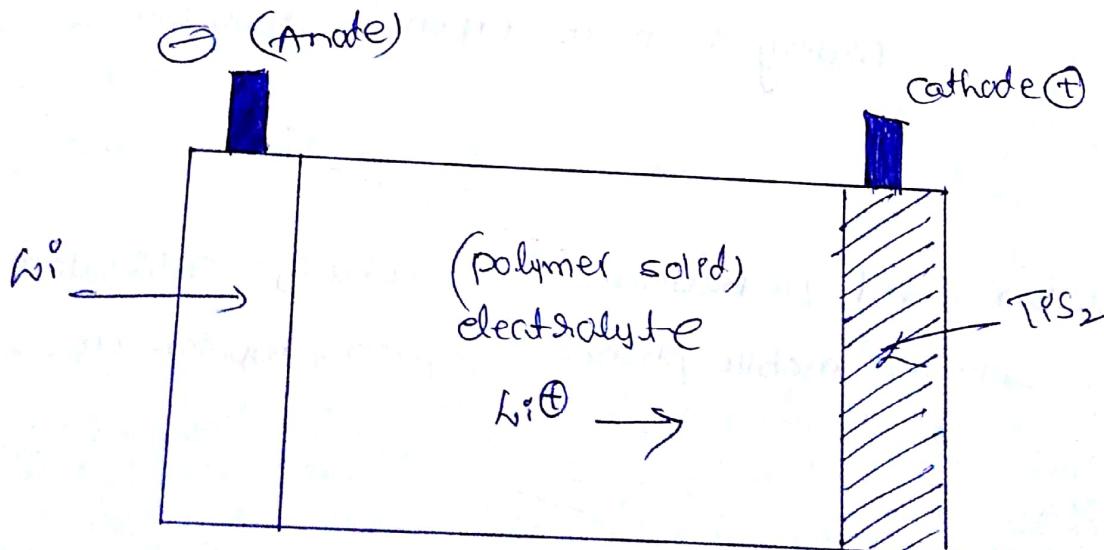
- A negative electrode (anode) (lithium metal)
- A positive electrode (cathode) (metal oxide or sulphide)
- An electrolyte (polymer)

### Construction:

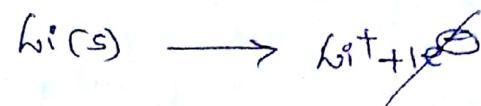
The lithium battery consists of a lithium anode and a  $TiS_2$  cathode. A solid electrolyte, generally a polymer is packed in between the electrodes. The electrolyte (polymer) permits the passage of ions but not that of electrons.

### Working [Discharging]

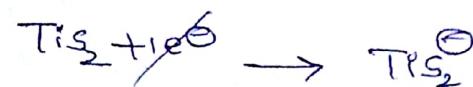
When the anode is connected to cathode, lithium ions move from anode to cathode. The anode is elemental lithium, which is the source of the lithium ions & electrons. The cathode is a material capable of receiving the lithium ions & electrons.



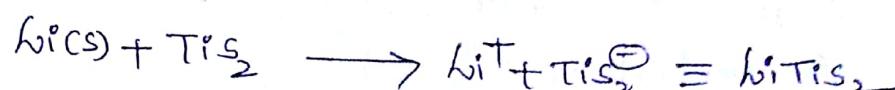
At anode:



At cathode:



Net chemical reaction is



### Advantages of Li Battery:

- (i) Its cell voltage is high 3.0V
- (ii) Since Li is a light-weight metal, only 7gms (1 mole) material is required to produce 1 mole of electrons.
- (iii) Since Li has the most negative  $E^\circ$  values, it generates a higher voltage than the other types of cells.
- (iv) Since all the constituents of the battery are solids there is no risk of leakage from the battery.

### Disadvantages of Li-Battery:

Li-Battery is more expensive than other batteries

### Uses:

Button sized Li Batteries are used in calculators, watches, cameras, mobile phones, laptop computers etc.

## Lead-acid storage battery:

A storage cell that can operate both as voltaic cell & as an electrolytic. When operating as a voltaic cell, it supplies electrical energy & as a result eventually becomes 'dendrite' when being recharged, the cell operates as an electrolytic cell.

This storage cell has the great advantage of working both as an electrolytic cell & as a voltaic cell.

Anode  $\rightarrow$  Lead (Pb)

Cathode  $\rightarrow$   $PbO_2$

Electrolyte  $\rightarrow$  dil  $H_2SO_4$  (20-21%)

Lead-acid storage battery consists of lead i.e. anode &  $PbO_2$  (lead dioxide) i.e. cathode. A number of Pb plates are connected in parallel & a number of  $PbO_2$  plates are connected in parallel. The Pb plates are fit in b/w the  $PbO_2$  plates, these plates are separated from adjacent one by insulators like strips of wood, glass fibre. The entire combination is immersed in dil  $H_2SO_4$  (electrolyte).

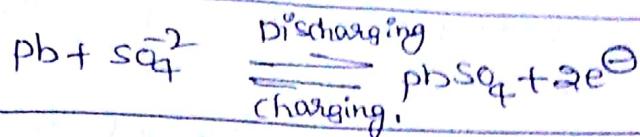
### → Discharging & charging reactions

The discharging of the storage cell is operating as a voltaic cell where the oxidation of lead takes place at anode.

At anode:

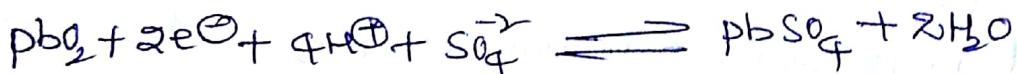
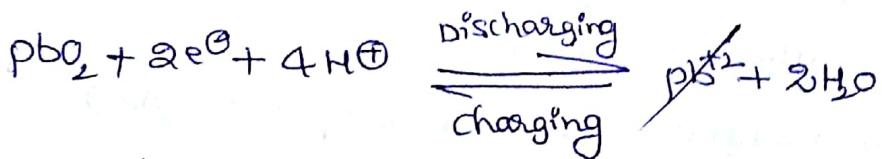


The  $\text{Pb}^{+2}$  ions are combined with  $\text{SO}_4^{2-}$  ions to produce  $\text{PbSO}_4$ .

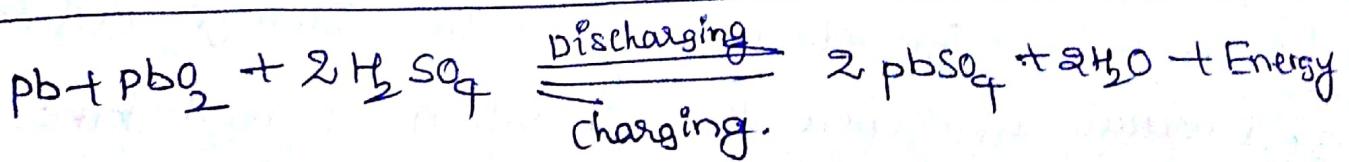
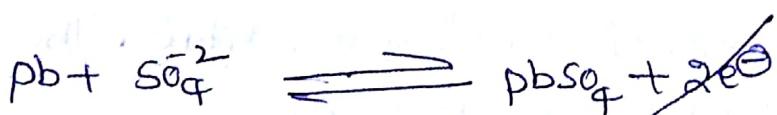


At cathode:

The  $2\text{e}^-$  are released at anode which are flows to the  $\text{PbO}_2$  electrode & causes reduction of  $\text{PbO}_2$  to produce  $\text{Pb}^{+2}$  which finally combines with  $\text{SO}_4^{2-}$  ions to produce at cathode.



Total net reaction is



The +ve pole of the generator is attached to the +ve pole of the battery & -ve pole of the generator is attached to the -ve pole of the battery. During discharging operation the concentration of  $H_2SO_4$  decreases, while the concentration of acid is restored during charging.

### Applications:

These cells are used to supply current for electrical vehicles, gas engine, in telephone exchanges, electric trains, mines etc.

### Lithium ion Battery:

Lithium-ion Battery is a secondary battery as in lithium cell, it doesn't contain metallic lithium as anode. As the name suggests, the movement of lithium ions are responsible for charging & discharging. Lithium-ion cell has the following three components.

- A positive electrode [layers of lithium-metal oxide (cathode)]
- A negative electrode [layers of porous carbon] anode
- An electrolyte (polymer gel)

### construction:

The positive electrode is typically made from a layers of chemical compound called lithium-cobalt oxide ( $LiCoO_2$ )

The negative electrode is made from layers of porous carbon (graphite)

Both the electrodes are dipped in a polymer gel electrolyte (organic solvent) and separated by a separator which is a perforated plastic and allows the lithium ions to pass through.

### Working:

#### Charging:

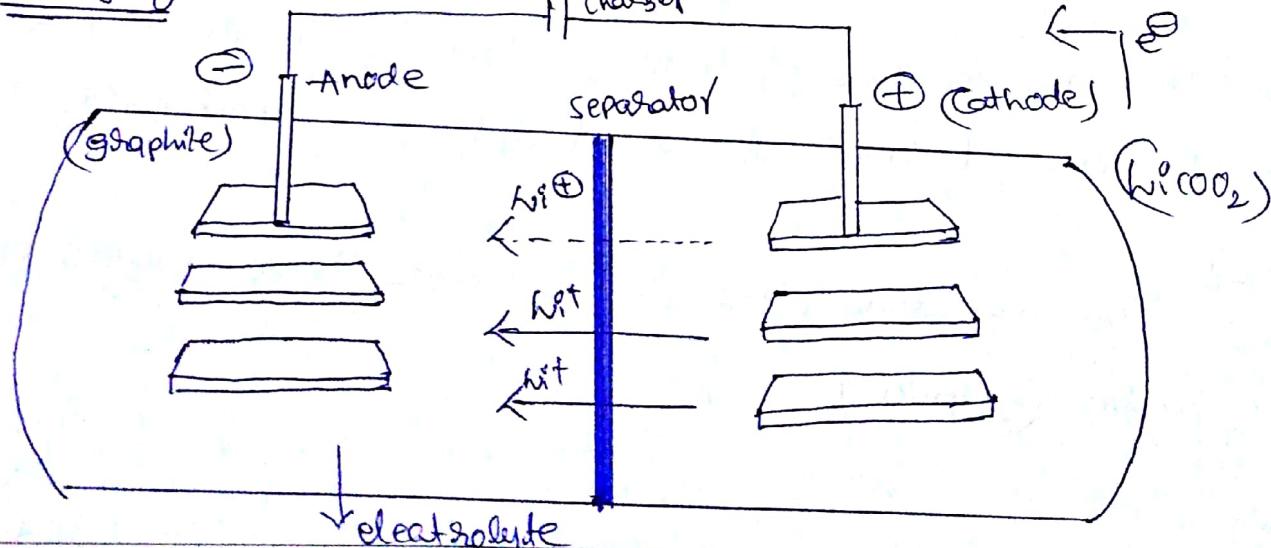
During charging, lithium ions flow from the positive electrode ( $\text{LiCoO}_2$ ) to the negative electrode (graphite) through the electrolyte.  $e^-$  also flows from the positive electrode to the negative electrode. The  $e^-$  & lithium ions combine at the negative electrode and deposit there as Li.



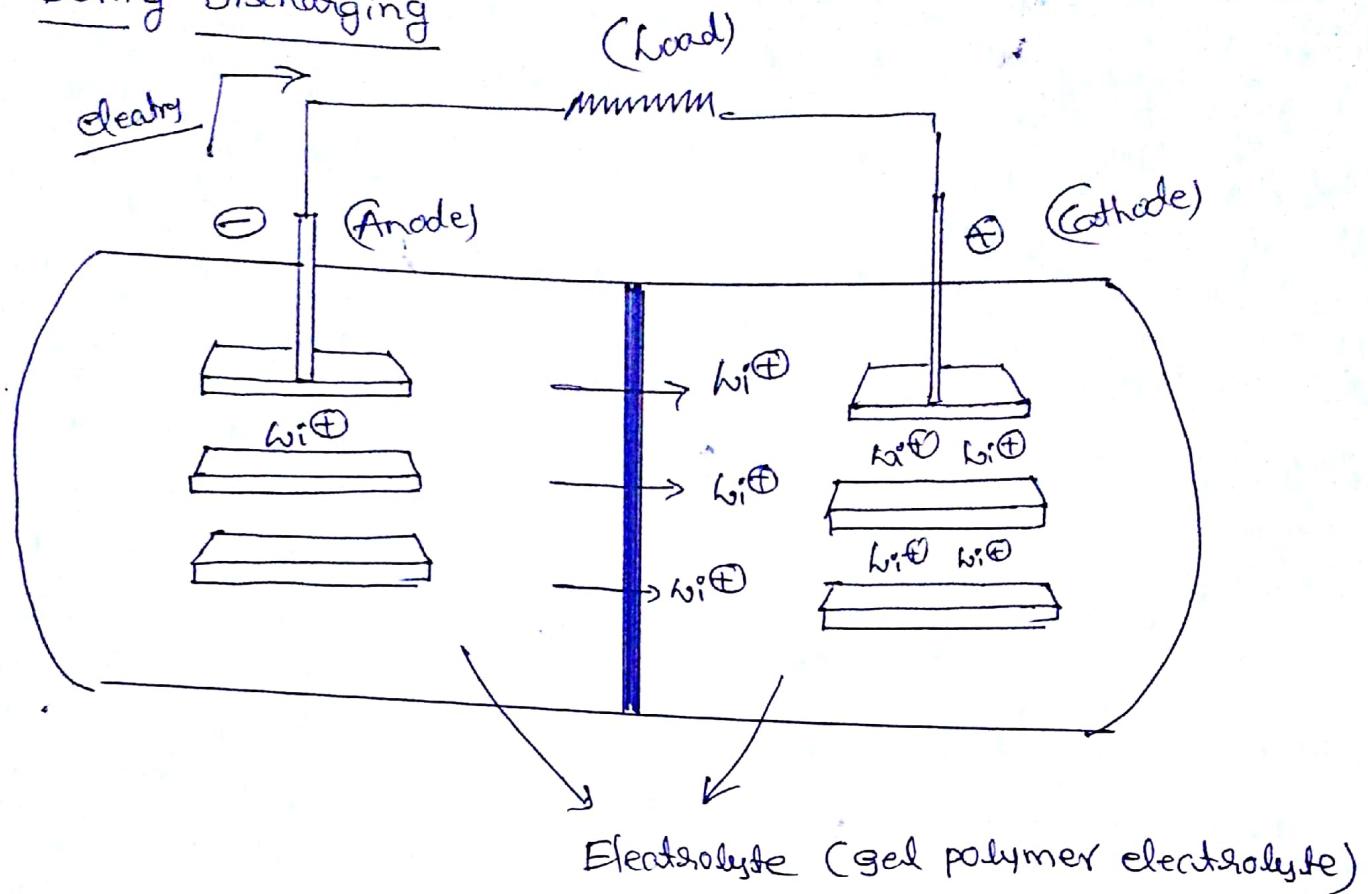
#### Discharging:

During discharging, the lithium ions flow back through the electrolyte from negative electrode to the positive electrode.  $e^-$  flows from the negative electrode to the positive electrode. The lithium ions and  $e^-$  combine at the positive electrode & deposit there as Li.

During Charging  $\rightarrow \text{Li}_{1-x}\text{COO}_2 + \text{C Li}_x \xrightarrow{\text{charger}} \text{LiCoO}_2 + \text{C}$



During Discharging



### Advantages:

- 1) Lithium ion batteries are high voltage and light weight batteries.
- 2) It is smaller in size.
- 3) It produces three times the voltage of Ni-Cd batteries.
- 4) It has none of the memory effect seen in Ni-Cd battery.

### Uses:

It is used in the cell phone, laptop, portable LCD TV, Semiconductor drive audio etc.

JGJW

## Introduction & scope:

Combined forms as oxides, hydroxides, carbonates, chlorides, sulphides, silicates etc are generally known as ores these are thermodynamically stable & these are reduced into their metallic states during the extraction process.

"The process of decay of metal by environmental attack is called corrosion." For example formation of reddish brown layer of rust on the surface of iron.

The corrosion of metal is measured in the units of milli inches/year (or) mm/year.

### Definition:

"The corrosion may be defined as the destruction of solid metallic material by chemical, electrochemical interactions with the environment."

When the metals are exposed to the environment the destruction of the metal starts at the surface of the metal due to the

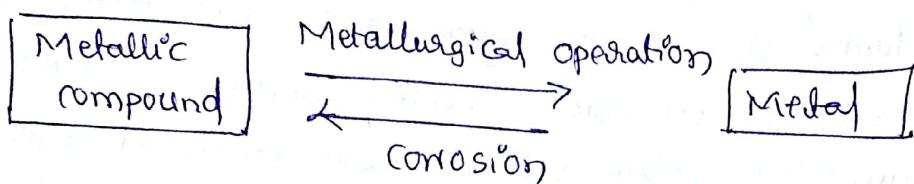
- 1) Chemical Reaction
- 2) Electro chemical Reaction.

### Causes:

The basic causes of corrosion is the instability of metals in their refined form.

Eg: Fe is exposed to the environment resulting in the formation of  $Fe_2O_3 \cdot nH_2O$  (Ferric oxide) (Or) Chemical composition of rust.

Generally the corrosion is a slow process & this corrosion takes place at the surface of the metals by producing a thin oxide layer.



### Effects of corrosion:

Effect of corrosion are briefly given below.

- 1) Loss of useful properties of metal & thus loss of efficiency.
- 2) Decrease in production rate, because efficiency is less & replacement of corroded equipment is time consuming
- 3) Increase in maintenance & production cost & contamination of product.

## Theories of corrosion:

Theories of corrosion are 2 types.

- 1) Chemical corrosion (or) dry corrosion
- 2) Electrochemical corrosion (or) wet corrosion.

### Chemical corrosion

This type of corrosion occurs mainly through the direct chemical action of environment (or) atmospheric gases such as Oxygen, halogen,  $H_2S$ ,  $SO_2$ ,  $N_2O$ , anhydrous inorganic liquid, these gases interact with the metallic surface & produce corrosion products.

There are three main type of chemical corrosion

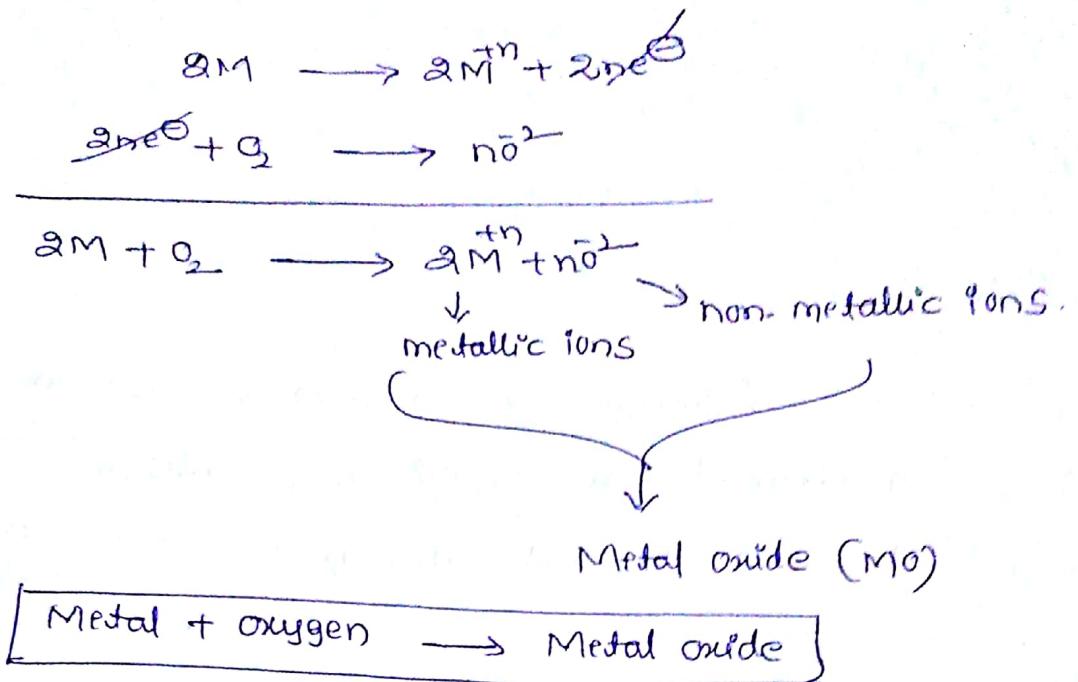
- 1) Oxidation corrosion
- 2) Corrosion by other gases
- 3) Liquid metal corrosion.

#### 1) Oxidation Corrosion

~~Ques~~  
~~Ans~~

"Direct attack of oxygen at high (or) low temp on metals in the absence of moisture is called oxidation corrosion". Generally metals are oxidised to a smaller extent however alkali metals ( $Li, Na, K, Rb$  etc) & alkaline earth metals ( $Ba, Ca, Sr$  etc) are very rapidly oxidised at low temperature except  $Ag, Au, Pt$  all metals are oxidised at high temperatures.

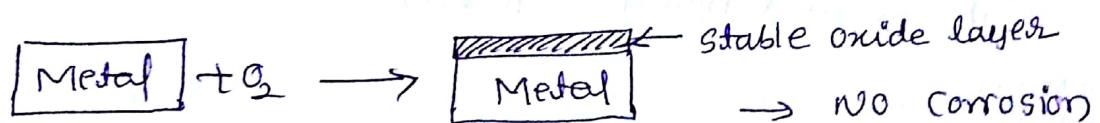
During the oxidation corrosion some of chemical reactions are takes place. They are.



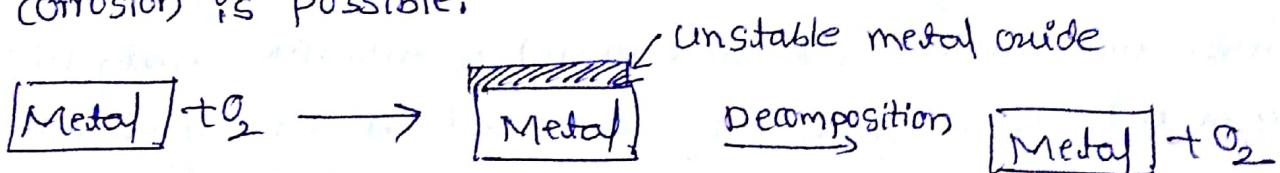
→ Nature of the metal oxide formed plays an imp role in oxidation corrosion.

- 1) A stable metal oxide layer with fine grained structure tightly adheres to the metal surface is impermeable to the attacking agent i.e oxygen. Such film behaves as protective coating in nature, thereby no corrosion takes place on the metal.

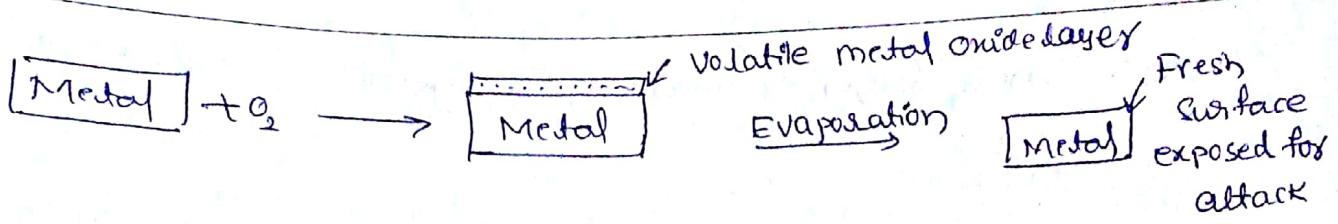
Ex: The oxide film on Al, Sn, Pb, Cu, Cr, Ni etc.



- 2) An unstable oxide formed on the metal surface decomposes back into the metal & oxygen. In such cases oxidation corrosion is possible.



- 3) A volatile oxide layer formed during corrosion evaporates as soon as it is formed, leaving the metal for further attack, this causes rapid & continuous corrosion leading to excessive corrosion to occur



To express the extent of protection given by the corrosion to the underlying metal Pilling-Bedworth rule was postulated.

Pilling-Bedworth ratio defined as 'the ratio of the volume of the metal oxide layer to the volume of the metal.'

$$P.B.W.R = \frac{\text{Volume of the oxide layer formed}}{\text{volume of the metal}}$$

- If the volume of the oxide layer is greater than the volume of the metal, then the layer is working as protective
- If the volume of the oxide layer is less than the volume of the metal, then the layer is working as further corrosion.

## 2) Corrosion by other gases:

When the metal is exposed to the environmental gases, like as  $S_2$ ,  $Cl_2$ ,  $H_2S$ ,  $F_2$ ,  $CO_2$  etc are interactants with the metal & the corrosion takes place at the surface of the metal.

The extent of the corrosion depends upon the chemical affinity in b/w metallic surface & gas involved.

- 1) If the film formed is non porous (or) protective then the corrosion is limited to only external surface

Eg:  $AgCl$  is formed by attack of  $Cl_2$  gas on  $Ag$ .

- 2) If the film formed is porous (or) volatile then the further corrosion is takes place.

Eg: Sncl<sub>4</sub> layer formed by attack of Cl<sub>2</sub> gas on Sn.

### Liquid Metal corrosion:

Liquid metal like Hg reacts with the metal & form alloys at higher temperature. This is caused because of the dissolution of a solid metal by the liquid metal.

### Electrochemical Corrosion (or) Wet corrosion

Wet (or) moist conditions through the formation of short circuited electrochemical cells, wet corrosion is more than dry corrosion.

All the metal being transformed from the metallic state to the ionic state. To complete the electrochemical reaction there must be two electrodes i.e. anode & cathode, there must be some potential difference is developed b/w the two electrodes which is used for the flow of es to the circuit.

Electrochemical corrosion can be explained on the basis of these electrochemical theory.

- Electrochemical corrosion occurs in the presence of eq/ solution.
- It occurs through a large no. of galvanic cells.
- This corrosion can be explained by the electrochemical reactions.
- Corrosion products (rust) are generally deposited on the cathodic area, further corrosion occurred.
- This corrosion is a very rapid process.

→ It occurs only on heterogeneous surfaces.

#### NOTE:

In chemical corrosion, corrosion products are deposited on the same spot where the corrosion starts hence further corrosion is prevented.

### Mechanism of Met (or) Electrochemical Corrosion

Electrochemical corrosion involves flow of es i.e current b/w the anodic & cathodic areas. The anodic reaction (i.e anodic area) involves in dissolution of metal as corresponding metallic ions with the liberation of free es.

#### At Anodic Area:



On the other hand, the cathodic reaction consumes the es with either by

- 1) Evaluation of hydrogen
- 2) Absorption of oxygen.

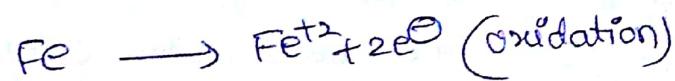
depending on the nature of the corrosive environment.

#### 1) Evaluation of hydrogen type

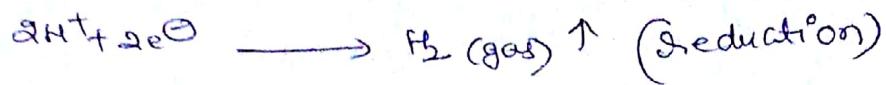
This type of corrosion occurs usually in acidic environments the rusting of iron takes place in acidic environment in the following way.

#### At Anode:

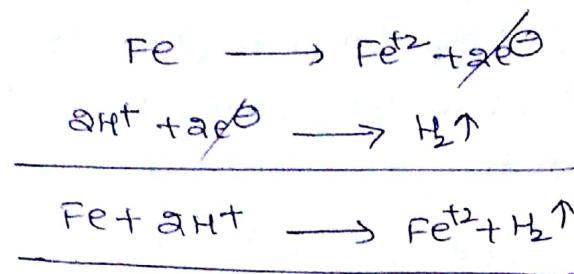
The dissolution (oxidation) of iron to  $Fe^{2+}$  ions with the liberation of electrons takes place.



These es flow through the metal, from anode to cathode where H<sup>+</sup> ions are available & eliminated as  $H_2$  gas ↑



The overall reaction is



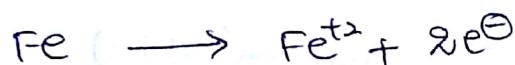
All metals above hydrogen in the electrochemical series have a tendency to get dissolved in acidic solution with simultaneous evolution of hydrogen.

### Absorption of oxygen:

Rusting of iron in neutral, aqueous solution of electrolyte like (NaCl solution) in the presence of  $O_2$  & some cracks developed in iron oxide film causes this type of electrochemical corrosion. The surface of iron is always coated with a thin film of iron-oxide. If a crack is developed in this iron oxide film, the anodic areas are created on the surface, while the well coated metal parts acts as cathode. It follows that the anodic areas are small & the cathodic areas are large.

#### At the Anodic area:

The metal oxidises as ferrous ions liberating  $e^-$ s which pass through metal to the cathodic area.

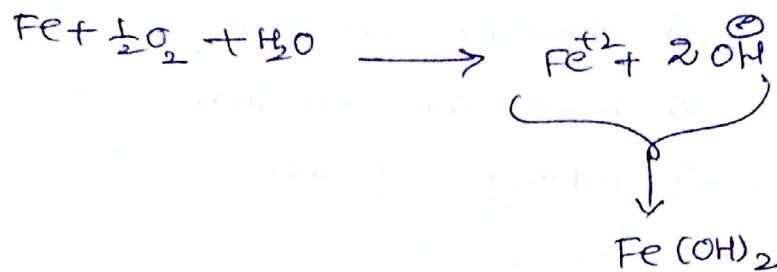
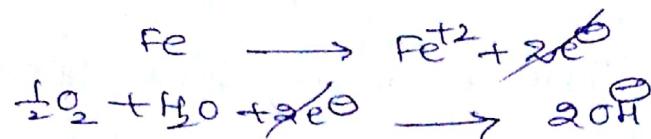


#### At Cathodic area:

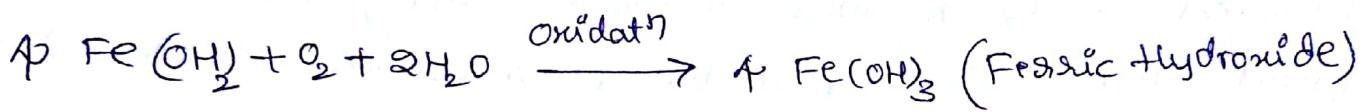
The  $e^-$ s were accepted by water & oxygen to generate  $OH^-$  ions by reduction.



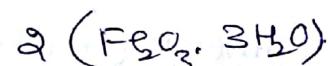
Overall reaction is



If enough  $\text{O}_2$  is present  $\text{Fe(OH)}_2$  is easily oxidised to ferric hydroxide & then to hydrated ferric oxide which is nothing but rust.



$\downarrow$   
oxidation.



rust.

J.S.S.

## Types of corrosion:

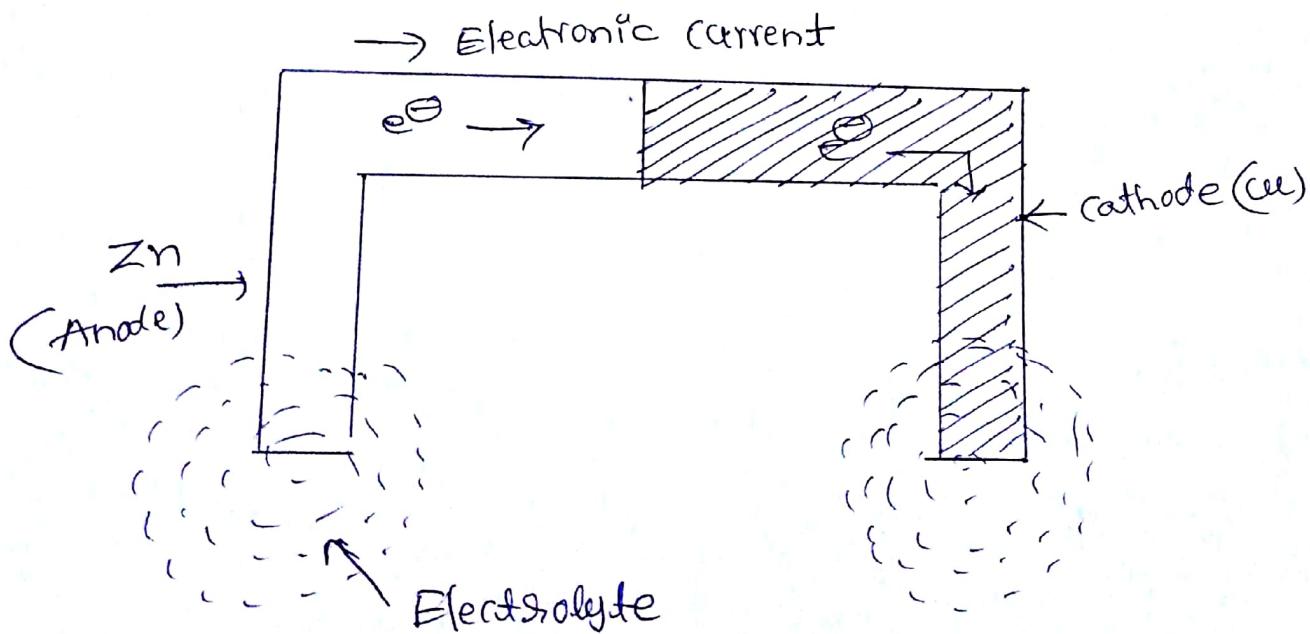
There are mainly three types of corrosion takes place they are

- 1) Galvanic corrosion
- 2) Metal line corrosion
- 3) Pitting corrosion.

### Galvanic corrosion:

When two dissimilar metals are connected & jointly exposed to the corrosive atmosphere, the metal possessing higher oxidation potential (or) higher in the electrochemical series becomes anodic & undergoes corrosion. The metal which is lower up in the electrochemical series it acts as a cathode.

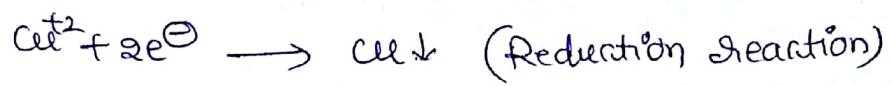
Ex: When Zn & Cu metals are connected & exposed to corroding environment (containing moisture, which acts as an electrolyte), Zn becomes anodic because of its higher oxidation potential so Zn undergoes oxidation & corroded whereas Cu undergoes reduction & protected.



At anodic metal



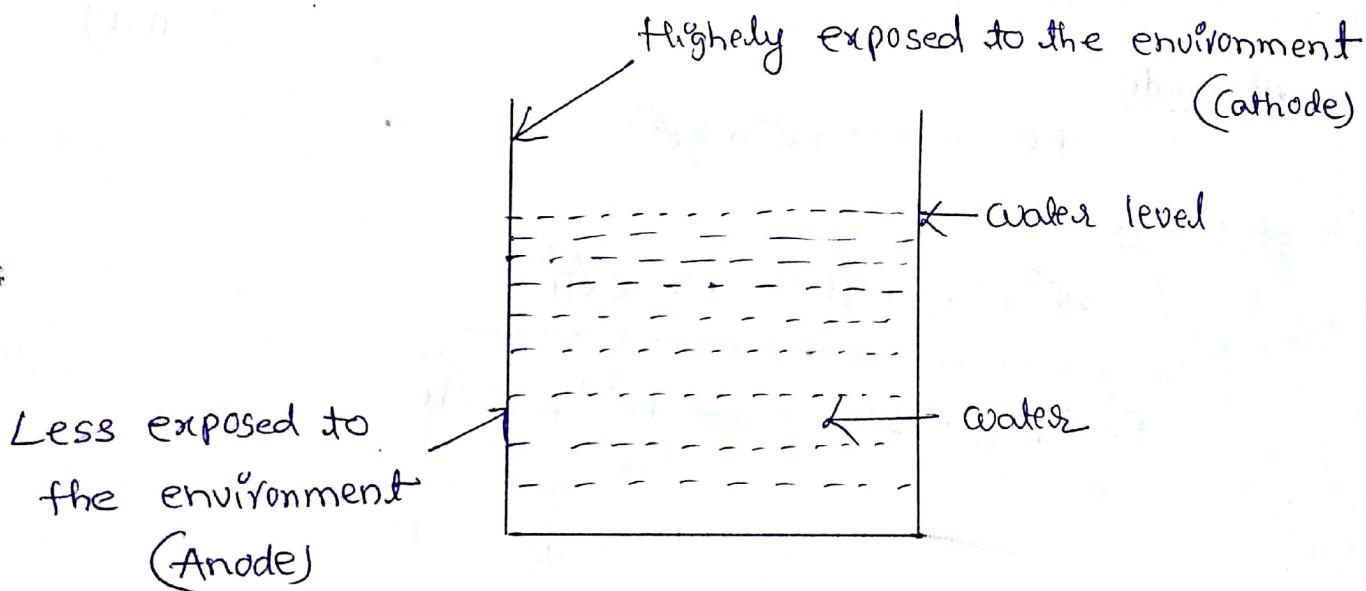
At cathodic metal



It is evident that corrosion occurs to metal anodic metal while the cathodic metal is protected.

When two (or) more than metals are used in a help of galvanic series. In an equipment using Zn in contact with Fe parts, Zn will be corroded because Zn is anodic to Fe when Fe is in contact with the Cu, Fe will be undergo corrosion because Fe is anodic to Cu.

### Waterline corrosion



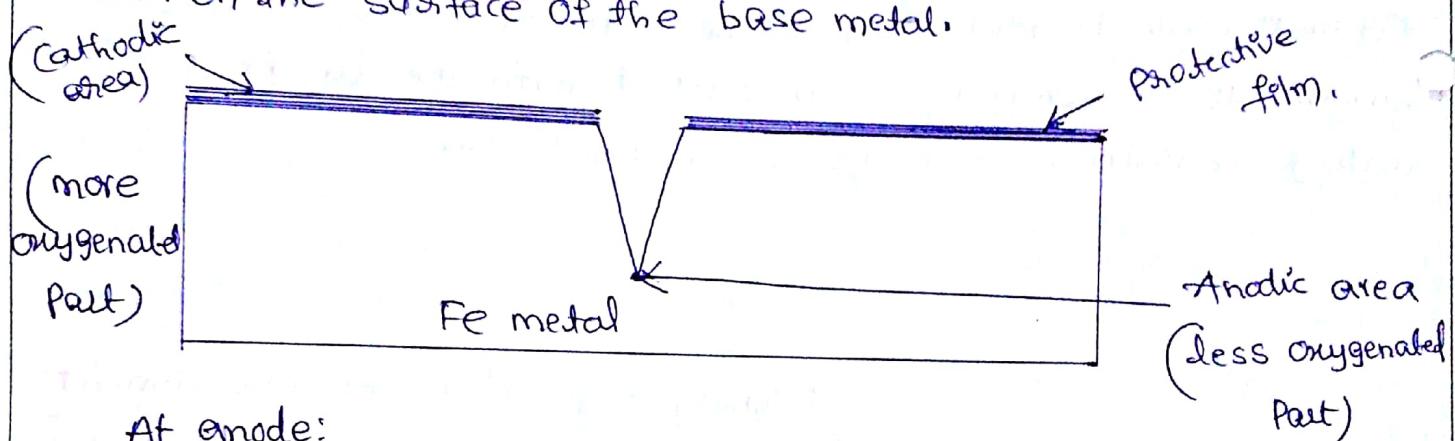
The concentration of  $\text{O}_2$  above the water level is higher than the below of the water level when water is taken in a steel tank. This is developed an oxygen concentration cell. In this cell the metal which above the water level is cathodic & the metal below the water level is anodic i.e one galvanic cell is formed in b/w the metal &

environment.

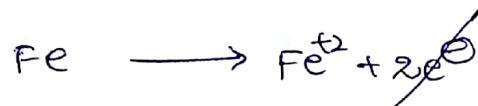
According to the principle of electrochemical corrosion i.e. corrosion occurs at anode & corrosion product formed at the surface of the cathodic metal.

### Pitting Corrosion:

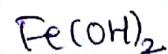
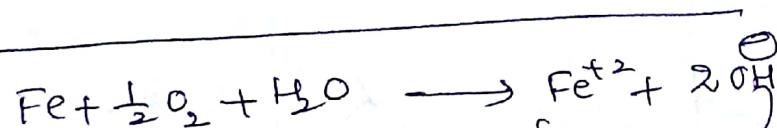
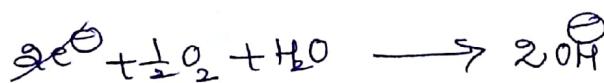
Any breaking & cracking of the protective film (coating layer) developed on the metal surface creates the formation of small anodic areas in the less oxygenated part & large cathodic areas in more oxygenated parts i.e. a pit is formed on the surface of the base metal.



#### At anode:



#### At cathode:



↓ Oxidation takes place.



#### NOTE:

one small pit is formed,  
The rate of corrosion will be increased.

## Factors effecting on corrosion:

The rate and extent of corrosion mainly depends up on the following factors.

- 1) Nature of metal
- 2) Nature of environment.

### Nature of metal.

#### (i) Metal position in the electrochemical series:

When two metals (or) alloys are in electrically contact in the presence of electrolyte, the more active metal (or) the metal having higher oxidation potential suffers from corrosion. The extent of corrosion depends on the position of the metal in the EMF series.

#### (ii) Purity of the metal:

When the metal is 100% pure, the metal will not undergo any type of corrosion. The metal which contains the impurities which form tiny electrochemical cells at the exposed parts of the metal. The extent of corrosion increases with increasing impurities.

% of purity of metal	99.999	99.99	99.93
corrosion rate	1	2650	500

Purity of metal  $\rightarrow$   $\downarrow$ ces the corrosion.

Impurity of metal  $\rightarrow$   $\uparrow$ ces the corrosion.

### (iii) Nature of the Oxide film:

Nature of the oxide (on surface film which can be explained by the pilling - Bed coath ratio i.e. volume of the metal oxide layer to the volume of the metal consumed).

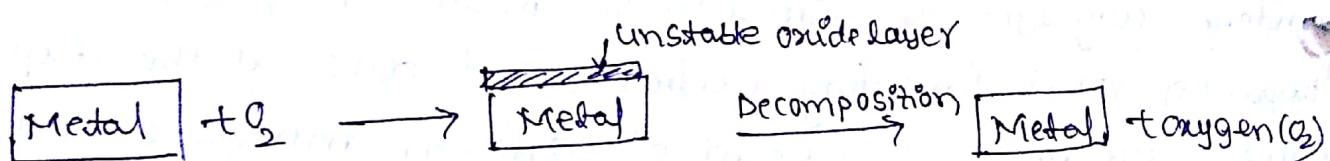
If ratio is higher than further corrosion doesn't take place because the surface of the metal which is completely covered by the oxide film which offering protect the metal from corrosion.

### (IV) Nature of the corrosion product:

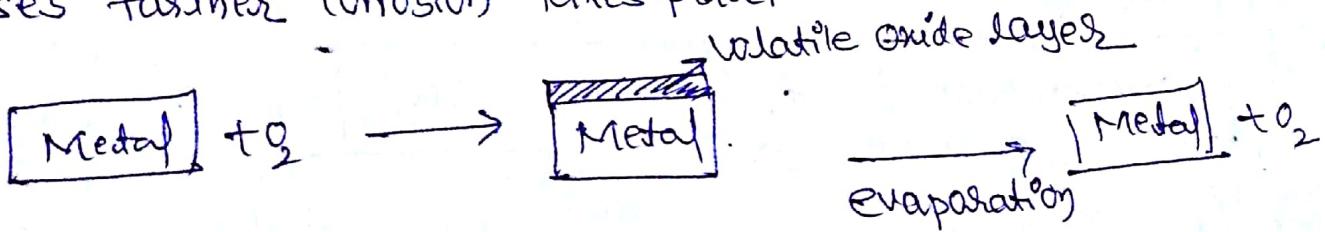
The state & extent of corrosion is mainly influenced by the following states of the corrosion products.

$\rightarrow$  If metal is form stable oxide layer so that layer acts as protective in nature i.e. resistant to further corrosion.

$\rightarrow$  If metal is form unstable oxide layer, so that layer decomposes back into the metal & oxygen. In such cases oxidation corrosion is possible i.e. further corrosion takes place.



$\rightarrow$  If the metal is form volatile oxide layer then it is as soon as it is evaporated back into the metal. In such cases further corrosion takes place.



## Nature of Environment:

### (i) Temperature:

If the temperature of environment increases the corrosion rate increases because the diffusion of ions are increases. Hence the

Corrosion rate increases with increasing temperature of environment  
corrosion rate decreases with decreasing temperature of environment.

i.e Rate of corrosion  $\propto$  Temperature.

### (ii) Effect of pH:

In generally acidic medium are ( $pH < 7$ ) more corrosive than alkali & neutral medium.

Eg: Zn is more (or) rapidly corroded in weakly acidic solution.

### (iii) Amount of oxygen in atmosphere:

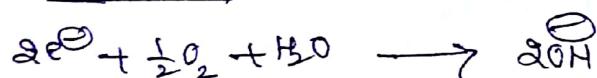
The rate of corrosion also increases due to the formation of oxygen concentration cell.

#### Anodic reaction:

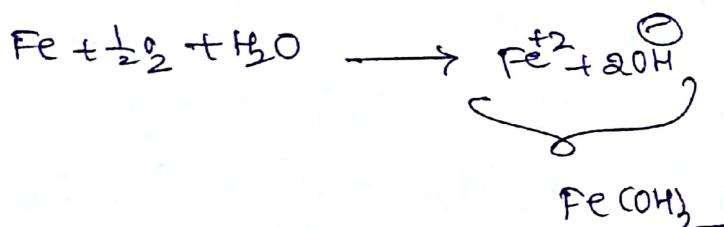


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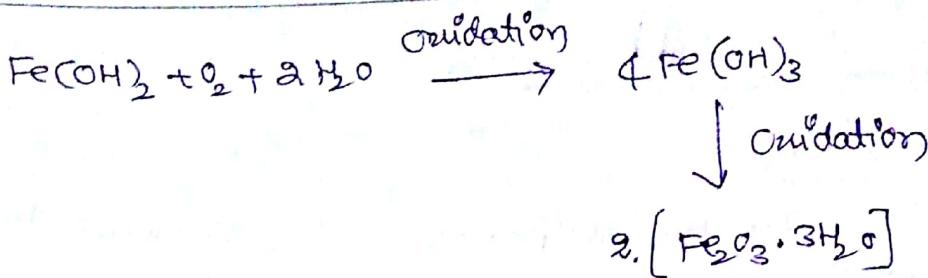
#### Cathodic reaction:



#### Overall reaction:



If enough oxygen present in atmosphere this  $\text{Fe(OH)}_2$  is converted into  $\text{Fe(OH)}_3$  (ferric hydroxide).



hydrated Ferric oxide (oxy)

Chemical composition of rust.

### Humidity:

The atmosphere contains the more moisture then the state & extent of corrosion is more, because of the moisture acts as a medium for oxygen in air & behaves as an electrolyte in air.

Eg: Atmospheric corrosion of Fe is slow in dry air when compared to Fe in moist air.



## Corrosion control methods:

As we have discussed the disadvantages & different mechanisms of corrosion so far, it is essential to know the different corrosion control methods. The following are the important control methods of corrosion.

- 1) proper designing
- 2) using pure metals & alloys.
- 3) modifying the environment
- 4) use of inhibitors
- 5) cathodic protection.
- 6) Application of protection coatings.

### Cathodic protection:

The method of protection given to a metal by forcibly making it to behave like a cathode is called cathodic protection. There are 2 types of cathodic protection.

- 1) Sacrificial anodic protection.
- 2) Impressed current cathodic protection.

#### (i) Sacrificial anodic protection:

In this method of protection, the metallic structure to be protected called "base metal" is connected to more anodic metal through a wire. The anodic metal undergoes corrosion slowly, while the base metal is protected. The corroded sacrificial anode block is replaced by a fresh one.

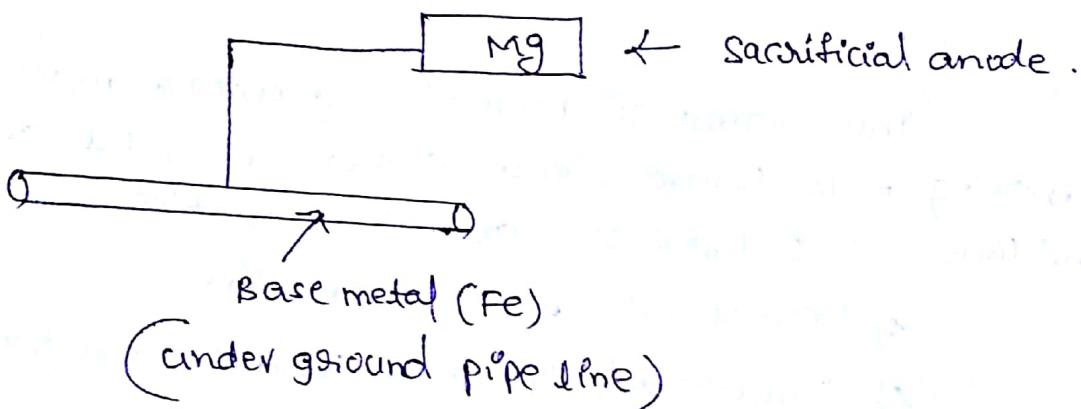
The commonly used anodic metals are Mg, Zn & Al & their alloys.

The important applications of sacrificial anodic protection are given below.

- 1) To protect marine structures & ship hulls, which are made of steel are connected to a sacrificial anode, which undergoes corrosion leaving the base metal protected.
- 2) Protection of buried pipelines, underground cables, water tanks, etc are also protected by sacrificial anode method.

By referring to electrochemical series, a small piece of the metal, anodic base metal is attached to the metal. The anodic metal undergoes corrosion & it's replaced from time to time

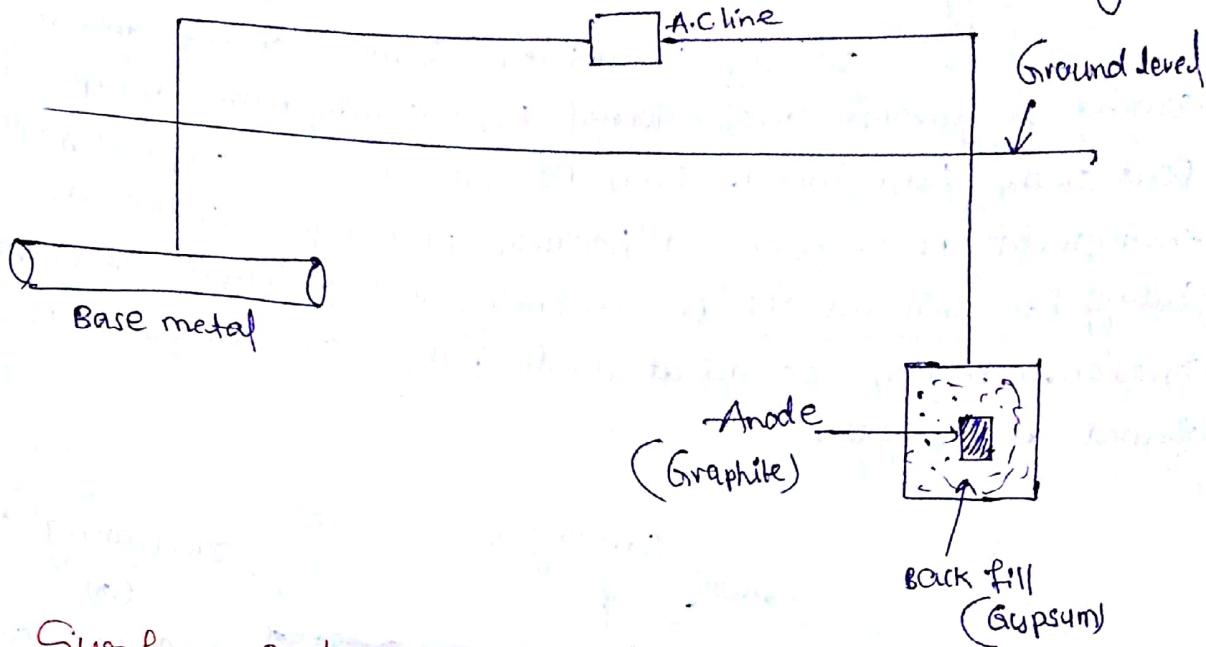
Ground level.



### Impressed Current Cathodic Protection:

In this method an impressed current little more than corrosion current is applied in the opposite direction, to nullify the corrosion current producing a reverse cell reaction thus the anodic corroding metal becomes cathodic & protected from corrosion. The impressed current is taken from a battery (or) rectified on A.C line. The anode is usually insoluble anode like graphite, high silica iron.

Usually a sufficient D.C current is passed onto the soluble anode kept in a "back fill" composed of coke (or) gypsum, so as to increase the electrical contact with the surrounding soil.



## Surface Coatings - Metallic Coatings

Protecting of the surface of the metal by applying the coating on the surface of the metal. This is common method to prevent the metal from corroding environment but the only limitation of this coating is coating which applied on the surface of the metal should be chemically inert to the environment at particular temperature & pressure.

Metallic Coatings are 2 types. They are

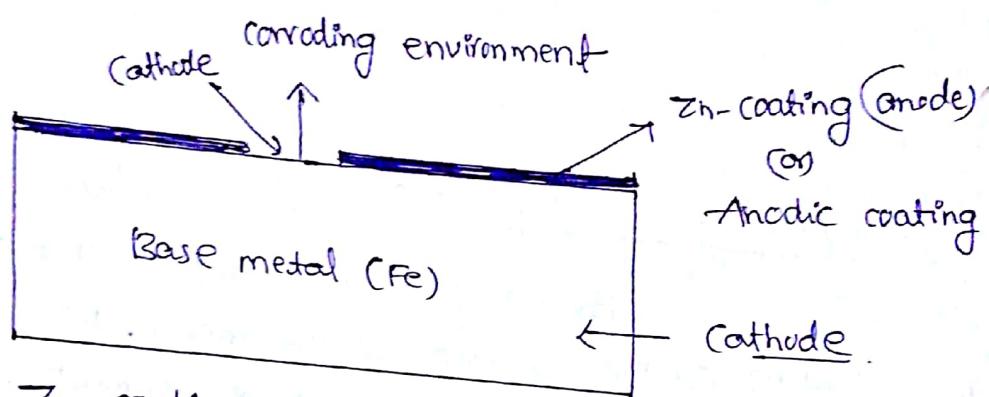
- 1) Anodic coatings.
- 2) Cathodic coatings.

### Anodic Coatings:

Metallic coatings are produced by coating one metal on the surface of another metal. The metal which is protected is called base metal & the metal which is coated on the surface of the base metal is called "coating metal".

This type of coatings are produced by coating an anodic coating metal on the surface of base metal.  
Eg: Coating of Zn, Al & Cd on Steel (or) Iron (Fe)

If any breaks are formed on the anodic part of the base metal, a galvanic cell is formed b/w coating metal & exposed environment. Even Fe (base metal) exposed to the environment. Fe is protected because when a part of the Zn coating is "scrapped off" i.e. according to the principle of galvanic cell corrosion, corrosion occurs at anode & the corrosion product is formed at cathode.



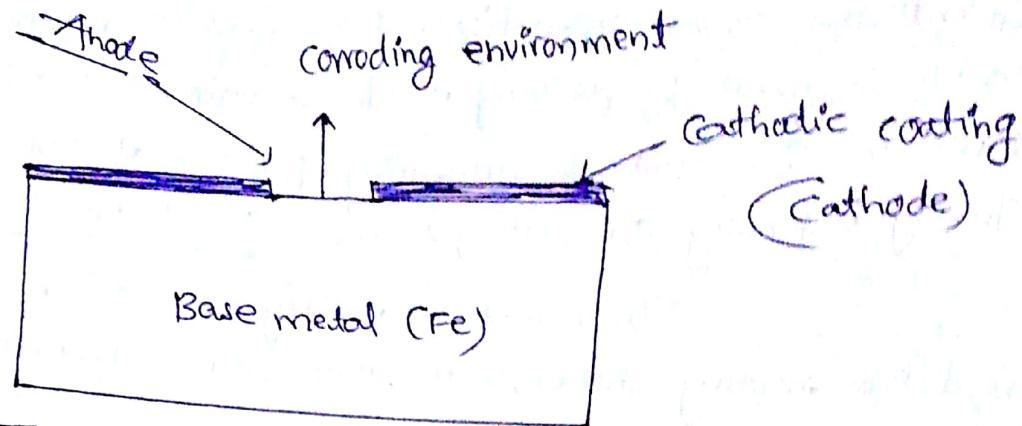
Eg: Zn-coating on Fe metal.

### Cathodic coating:

Cathodic coatings are obtained by coating a more noble metal (having higher reduction potential) than the base metal. Because the cathodic metal has high corrosive resistance character i.e. cathodic coatings are acts as high corrosive resistance to the base metal when coating is completely continuous & free from breakage.

If any breakage occurs on the cathodic coating than the further corrosion takes place. When the base metal which is anodic to the coating metal is exposed to environment, a galvanic cell is setup resulting severe pitting & perforation in the base metal because of a small anode & large

Cathodic area.



Eg: Tin coating on Fe metal.

### Methods of application on metal

Of the metallic coatings. The following are the methods of application

#### Hot dipping:

Such as Zn, Sn, Pb, Al etc on Iron metal which have higher melting points. This is a method of coating a low melting metal coating metal which is covered by a molten flux layer which cleans the surface of the base metal & prevents the oxidation of the coating metal. For good adhesion of the coating metal on the surface of base metal, the base metal surface must be very clean.

The most widely used hot dipping methods are

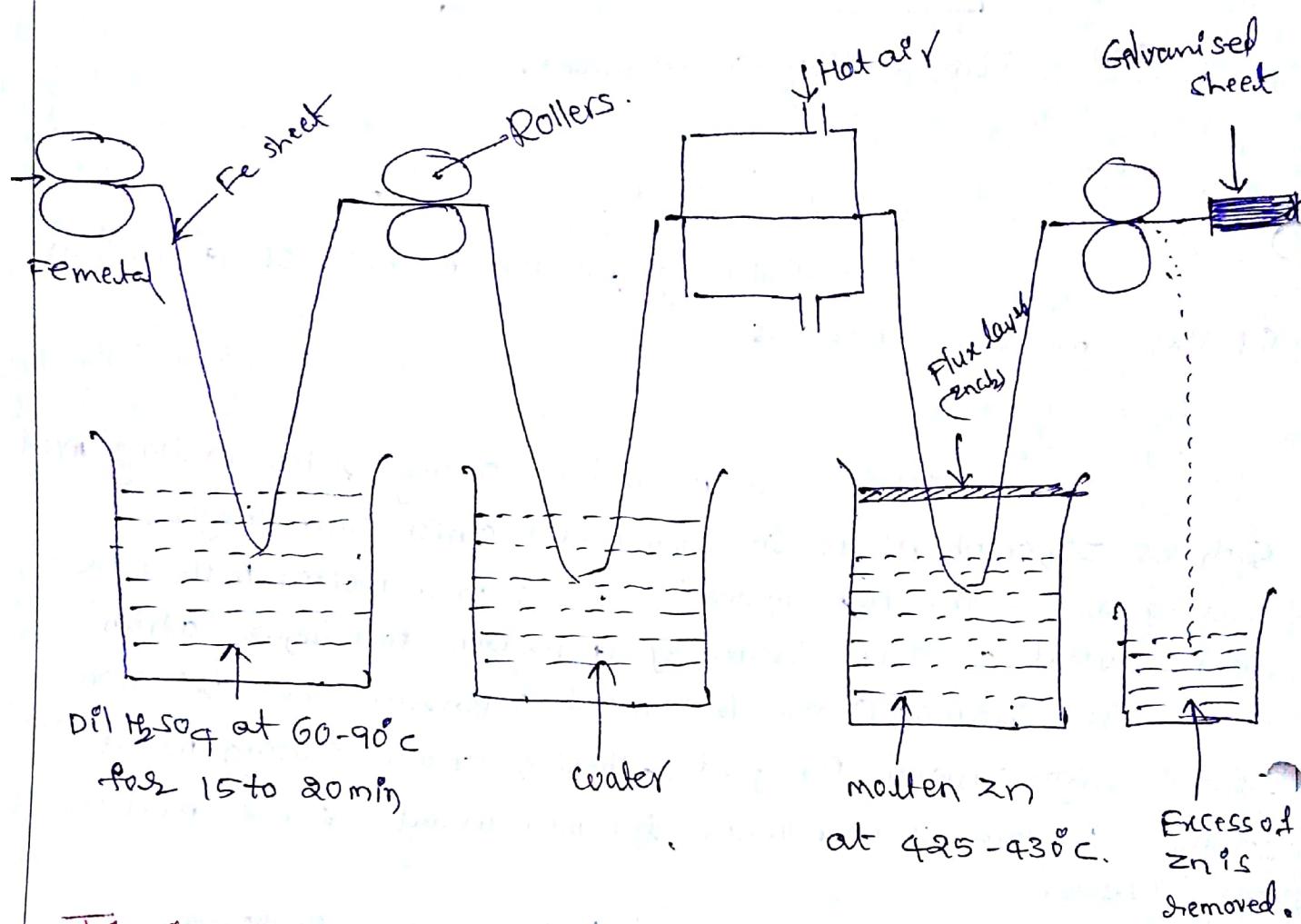
- 1) Galvanization
- 2) Tinning.

#### Galvanization:

It is a process of coating Fe sheet (or) steel is cleaned by acid pickling method with dil  $H_2SO_4$  for 15-20 min at 60-90°C. The sheet is then washed well & dried it is dipped

in a bath of molten Zn maintained at  $425-430^{\circ}\text{C}$ . The surface of the bath is kept covered with flux layer to prevent the oxide formation. The sheet is taken out & excess of Zn is removed by passing it b/w a pair of hot rollers. Then the sheet is subjected to annealing process at  $650^{\circ}\text{C}$  & cooled slowly & finally we will get the coated Fe sheet by Zn metal.

This method mostly used to protect iron(Fe) used for roofing sheets, wires, pipes, nails, bolts etc.



### Tinning:

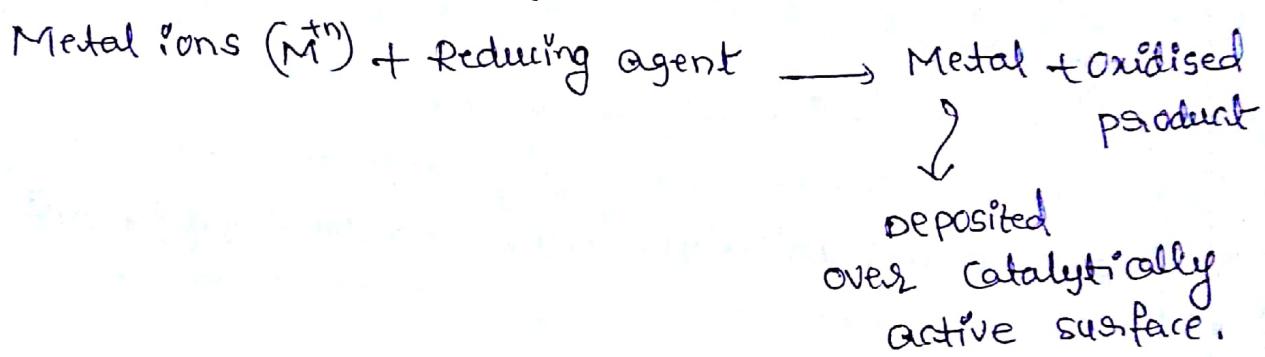
It is a process of coating of Fe sheet (or) steel sheet by tin metal. In this process surface of the base metal i.e. Fe sheet is cleaned by acidic pickling with dil  $\text{H}_2\text{SO}_4$  & passed through a bath of  $\text{ZnCl}_2$  flux. The flux help the molten metal to adhere to the metal surface. Then the sheet is passed through molten tin bath & passed b/w two rollers. Then we will get the

Coated Fe sheet by tin metal.

### Electroless plating:

Electroless plating is a technique of depositing a noble metal (from its salt solution) on a catalytically active surface of a less noble metal by employing a suitable reducing agent without using electrical energy called electroless plating.

The added reducing agents, causes the reduction of the metallic ions to the metal, which eventually gets plated over the catalytically activated surface.



### Ex: Electroless Ni plating:

It involves the following features  
pretreatment & activation of the surface.

The surface to be plated is first degreased by using organic solvents (or alkali), followed by acid treatment.

(a) Metals & alloys like Al, Cu, Fe, brass etc be directly nickel plated without any activation.

- (b) stainless steel is activated by dipping in hot solution of 50% dilute  $H_2SO_4$ .
- (c) Activation of Mg alloy surface is carried out by giving a thin coating of Zn & Cu over it

### Composition of bath:

Coating solution  $NiCl_2$  solution — 20 gm/lit

Reducing agent — Sodium hypo phosphite

buffer — sodium acetate (20 gm/lit)

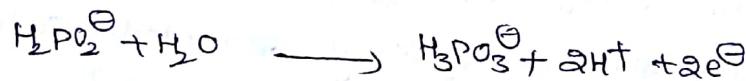
Complexing agent — (10 gm/lit)

Optimum pH — 4.5

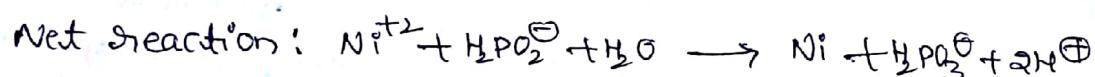
Optimum temp — 93°C.

### Reactions:

#### At Anode:



#### At Cathode:





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# **UNIT-III**

# **QUESTION BANK**

**OR**

**UNIVERISTY**

**QUESTIONS**



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### UNIT-III

## ELECTRO CHEMISTRY & BATTERIES, CORROSION

### JNTUH PREVIOUS QUESTIONS

1. (a) Outline the electrochemical theory of corrosion and explain the mechanism of rust formation in acidic and neutral medium.  
(b) Write a brief account of cathodic protection. (DEC-2010)
2. (a) Write a short note on electro chemical series?  
(b) Explain the functioning of a Galvanic cell? (DEC-2010)
3. (a) Define the terms specific, equivalent and molar conductivities. How do they vary with dilution.  
(b) Calculate the cell constant of a cell having a solution of Concentration/30gm. equiv/litre of an electrolyte which showed the Equivalent conductance of 120 Mhos cm<sup>2</sup> gm equiv(DEC-2010)
4. (a) Discuss the influence of following factors on corrosion:-
  - i. Over voltage
  - ii. Nature of the metal
  - iii. Nature of environment.(DEC-2010)
5. Explain the nature and role of constituents of organic paints. (DEC-2010)
6. Distinguish Electrochemical cells from Electrolytic cells with suitable illustrations. (DEC-2010)
7. What is electrode potential? How is it determined by using Calomel Standard electrode. (DEC-2010)
8. Explain sacrificial anodic protection method of controlling corrosion. (DEC-2010)
9. Write a note on anodic protection and the nature of corrosion product. (DEC-2010)
10. What is cell constant? How is it determined? (DEC-2010)
11. Describe a method for the determination of pH of a solution using Standard Calomel electrode. (DEC-2010)
12. Write notes on the following:-



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- (a) Hot dipping  
(b) Galvanizing  
(c) Tinning  
(d) Electroplating.
- 13.(a) What are concentration cells? How can the EMF of a concentration cell be evaluated? (DEC-2010)
14. Write short notes on single electrode potential and its significance.  
(a) Explain the process of galvanizing and tinning.  
(b) What are organic paints? Describe their constituents. (JUNE-2010)
15. What do you understand by electrochemical series? How is this series useful in the determination of corrosion of metals? (JUNE-2010)
16. The resistance of a 0.1N solution of an electrolyte of 40 ohms. If the distance between the electrodes is 1.2 cm and area of cross section is 2.4 cm<sup>2</sup>. Calculate equivalent conductivity. (JUNE-2010)
17. Discuss the influence of following factors on corrosion:-  
i. Over voltage  
ii. Nature of the metal  
iii. Nature of environment. (JUNE-2010) (DEC-2010)
18. Explain the nature and role of constituents of organic paints. (JUNE-2010)
19. Explain the electrochemical theory of corrosion of metals with special reference to rusting of iron. (JUNE-2010)
20. Write a note on galvanizing and metal cladding? (JUNE-2010)
21. Give reasons for the following statements:-  
i. When a zinc rod is dipped in a solution of aq. copper sulphate, copper is precipitated out.  
ii. Nernst equation is applicable for the determination of emf of a concentration cell. (JUNE-2010)
22. How are metals protected by impressed current method? (JUNE-2010)
23. Explain the galvanization and tinning processes of metals. (JUNE-2010)
- 24.(a) Define the terms specific, equivalent and molar conductivities. How do they vary with dilution?  
(b) Calculate the cell constant of a cell having a solution of concentration N/30gm. equiv/litre of an electrolyte which showed the equivalent



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- conductance of 120 Mhoscm<sup>2</sup> gm equiv $\square$ 1: (JUNE-2010)
- 25.What is a single electrode potential? Describe a method for its determination. (JUNE-2011)
- 26.Write the cell reactions of a Daniel cell. (JUNE-2011)
- 27.Write a brief account of concentration cells. (JUNE-2011)
- 28.Describe the different types of corrosion and discuss the factors that affect corrosion. (JUNE-2011)
- 29.What is cathodic protection? (JUNE-2011)
- 30.Write a note on electro plating. (JUNE-2011)
- 31.What is an electro chemical cell? Explain the construction and reactions of any electro chemical cell. (JUNE-2011)
- 32.Explain how Nernst equation is useful in calculating the electrode potential? (JUNE-2011)
- 33.Define metallic corrosion? Explain the mechanism of electrochemical corrosion by hydrogen evolution and oxygen absorption. (JUNE-2011)
- 34.Explain differential aeration corrosion with a suitable example.(JUNE-2011)
- 35.Discuss the phenomenon and applications of
- i. metal cladding and
  - ii. electroplating. (JUNE-2011)
- 36.Explain the principle of the hydrogen-oxygen fuel cells. (JUNE-2011)
- 37.Differentiate between electrolytic cells and concentration cells with suitable examples. (JUNE-2011)
- 38.Define the terms:
- i. Specific conductance
  - ii. Equivalent conductance of an electrolyte. How do they vary on dilution? (JUNE-2011) (JUNE-2011)
- 39.Explain the principle involved in wet corrosion. (JUNE-2011)
- 40.What are organic paints? Explain their constituents and functions.(JUNE-2011)
- 41.What are potentiometric titrations? Explain their applications. (DEC-2011)
- 42.Explain the corrosion of metals and the different types of corosions. (DEC-2011)



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43. Give an account of electrochemical theory of corrosion. (DEC-2011)
44. (a) How are metals protected by impressed current method? (DEC-2011)  
(b) Explain the galvanization and tinning processes of metals. (DEC-2011)
45. Write a note on single electrode potential and ion selective electrodes. (DEC-2011)
46. Explain the electrochemical theory of corrosion of metals with special reference to rusting of iron. (DEC-2011)
47. Write a note on galvanizing and metal cladding. (DEC-2011)
48. Describe the experimental method for the determination of the pH of a Solution, using quinhydrone electrode. (DEC-2011)
49. What are conduct metric titrations? Explain their applications. (DEC-2011)
50. Write a brief account on dry corrosion. Explain the factors affecting dry corrosion. (DEC-2011)
51. What is cathodic protection? How is it done by using sacrificial anode method? (DEC-2011)
52. The specific conductivity of a N/50 solution of NaCl at 300°C is 0.003686 ohm $\square$ 1cm $\square$ 1. If the resistance offered by the solution when placed in a cell is 1,500 ohms, calculate Cell constant and Equivalent conductance of solution. (DEC-2011)
53. Write a note on hydrogen-oxygen fuel cells?
54. Explain about the causes and effects of corrosion.
55. What is oxidation corrosion? How does it takes place? Explain about the mechanism of oxidation corrosion (JUNE-2013)