

Electro-chemistry

(1)

Electro chemistry : This is a branch of physical chemistry that deals with conversion of chemical energy ~~into~~ ^{heat} into electrical and vice versa.

Electrode : Electrode is a rod which is immersed in a electrolytic solution.

Electrolysis : The phenomenon of decomposition / break down/ lysis of electrolyte by passing electricity is called electrolysis.

Electrolyte : A substance which decomposes as a result of passing of electric current is called "Electrolyte".

Anode : It is an electrode at which oxidation occurs.

Cathode : It is an electrode at which Reduction occurs.

- Electrolytes may be classified in to two types.

(1) strong Electrolyte - Ex: HCl , H_2SO_4 , HNO_3 , NaOH , KOH , NaCl , KCl etc.

(2) weak electrolytes - Ex: Acetic acid, propionic acid, alkylamines, NH_4OH etc.

Electrode potential (E) :

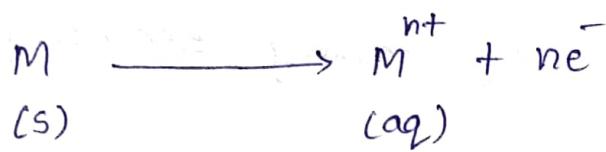
The tendency / capacity of metal is converted in to metal ion or metal ion is converted in to metal

is called as electrode potential of the cell (E_{cell}).

- the flow of current from one electrode to another electrode in galvanic cell indicates that the two electrodes have different potentials.

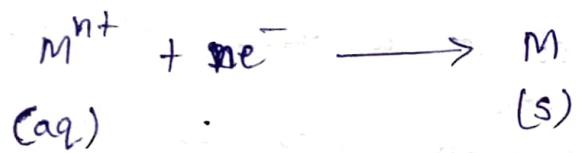
Oxidation potential (E_{Oxed} / E_L) :

The tendency / capacity of metal is converted into metal ions is called oxidation potential.



Reduction potential (E_{Red} / E_R) :

The tendency / capacity of metal ion is converted into metal is called as Reduction potential.



- reduction potential is always high when compared to oxidation potential.

$$E_{cell} = E_{Red} - E_{Oxed}$$

(or)

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

Standard Electrode potential (E°) : (SEP)

When this Electrode potential is measured under standard conditions like 25°C (Temp), 1.0 M concentration and pressure of 1 atm, then the electrode potential is called "standard electrode potential" (SEP).

- standard electrode potential of oxidation half cell is standard oxidation electrode potential (E_{Ox}°)/ (E_L°).
- standard electrode potential of Reduction potential half cell is standard reduction electrode potential (E_R°)/ (E_{Red}°).
- E_{Red}° is always high when compared to E_{Ox}° .

$$E_{\text{cell}}^\circ = E_{\text{Red}}^\circ - E_{\text{Ox}}^\circ$$

(or)

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

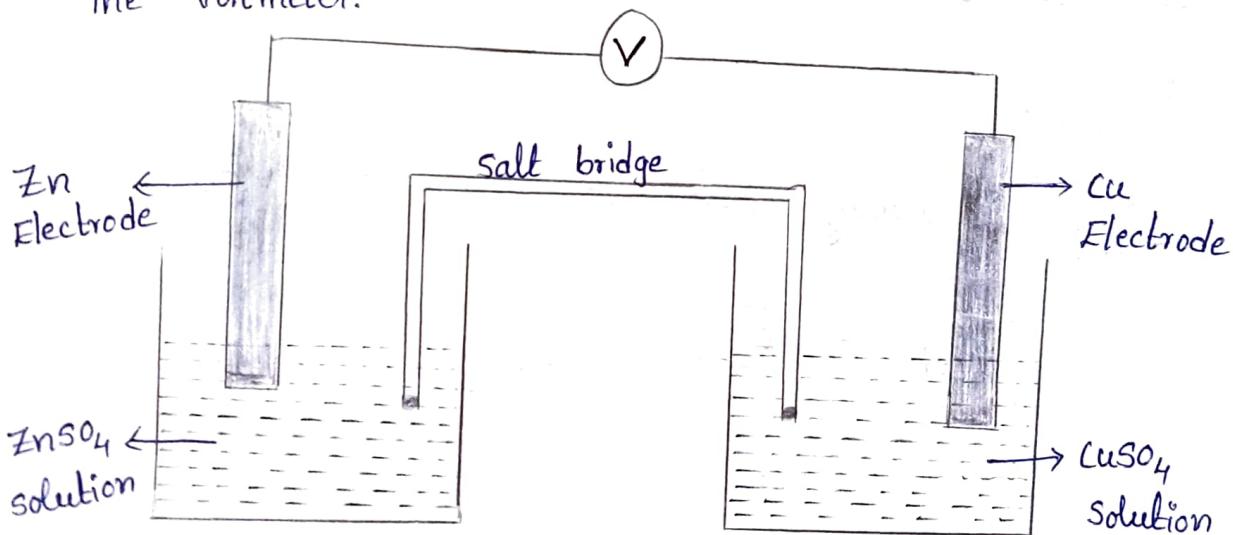
Electrochemical cell :

Galvanic cell / voltaic cell :

A voltaic cell is a device which is used for converting chemical energy into electrical energy.

Cell construction : (Ex: Daniel cell)

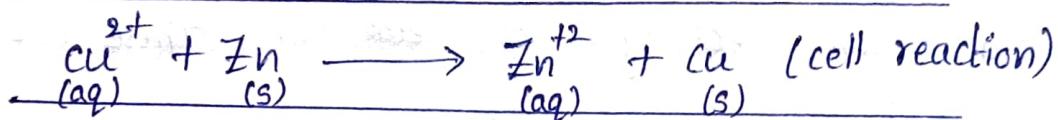
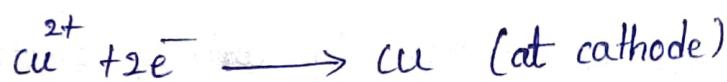
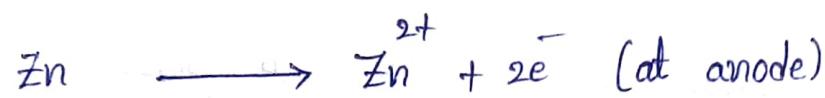
- Daniel cell consists of a Zinc electrode dipped in 1M ZnSO_4 solution and copper electrode dipped in 1M CuSO_4 solution. Each electrode is known as a half cell.
- The two solutions are interconnected by a salt bridge and the two electrodes are connected by a wire through the voltmeter.



Reactions occurring in the cell :

At Anode : oxidation takes place in the zinc electrode by the liberation of electrons.

At cathode : Reduction takes place in the copper electrode by the acceptance of electrons.



- The electrons liberated by the oxidation reaction flow through the external wire and are consumed by the copper ions at the cathode.

salt bridge :

It consists of a U-shaped tube containing NaNO_3 / NH_4NO_3 in agar-agar (solidifying agent). It connects to the two half cells of the galvanic cells.

- It is used for two maintain the equal potential of two cells.
- If oxidation potential is increases the salt bridge diffuse Na^+ / NH_4^+ ions, If reduction potential is increases the salt bridge diffuse NO_3^- ions to the half cells.

Representation of cell :

- i) A galvanic cell consists of two electrodes anode and cathode.

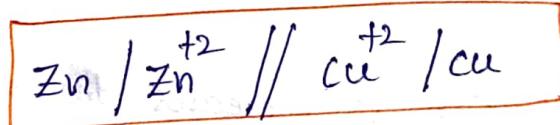
- (iii) The anode is written on the left hand side while the cathode is written on the right hand side.
- (iv) The anode must be written by writing electrode first and then electrolyte or metal ion. These are separated by vertical line.



- (iv) The cathode must be written by writing electrode ions or electrolyte first then electrode (metal). These are separated by vertical line.



- (v) The two half cells are separated by a salt bridge which is indicated by two vertical lines.



(or)



- The flow of electrons from one electrode to other electrode in a galvanic cell indicates that the electrodes have different potentials.

The difference of potentials which causes electrons to flow from one electrode which is at higher potential to another electrode which is at lower potential is called "Electro motive force (EMF)" of the cell.

Nernst Equation for EMF of a cell :

Any Electro chemical reactions,



Relative to hydrogen electrode is given by a simple form of the Nernst equation as

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

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

Here,

$$E_{M^{n+}/M} = E_{\text{Reduction}}$$

$$E_{M/M^{n+}} = E_{\text{oxidation}}$$

$E_{M^{n+}/M}^{\circ}$ = standard electrode potential for reduction. (This is determined by using the standard hydrogen electrode)

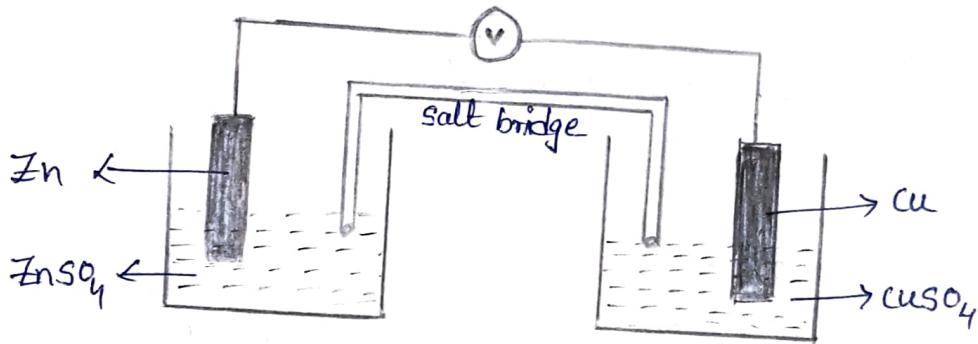
R = universal gas constant (8.31 kJ/mole)

T = Temperature in kelvin

n = No. of moles of electrons

F = Faraday (96500 C)

In Daniel cell,



The electrode potential of the right hand electrode is written as,

$$E_{\text{Cu}^{+2}/\text{Cu}} = E_{\text{Cu}^{+2}/\text{Cu}}^{\circ} + \frac{RT}{nF} \ln [\text{Cu}^{+2}(\text{aq})]$$

The electrode potential of left hand electrode is written as,

$$E_{\text{Zn}/\text{Zn}^{+2}} = E_{\text{Zn}/\text{Zn}^{+2}}^{\circ} + \frac{RT}{nF} \ln [\text{Zn}^{+2}(\text{aq})]$$

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

$$E_{\text{cell}} = \left(E_{\text{Cu}^{+2}/\text{Cu}}^{\circ} + \frac{RT}{nF} \ln [\text{Cu}^{+2}(\text{aq})] \right) - \left(E_{\text{Zn}^{+2}/\text{Zn}}^{\circ} + \frac{RT}{nF} \ln [\text{Zn}^{+2}] \right)$$

$$= \left[E_{\text{Cu}^{+2}/\text{Cu}}^{\circ} - E_{\text{Zn}^{+2}/\text{Zn}}^{\circ} \right] + \frac{RT}{nF} \ln \frac{[\text{Cu}^{+2}]}{[\text{Zn}^{+2}]}$$

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

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

By converting natural logarithm to the base 10, and substituting the values R, T (298 K) and F,

We get,

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} + \frac{0.0591}{2} \log \frac{[\text{Cu}^{+2}(\text{aq})]}{[\text{Zn}^{+2}(\text{aq})]}$$

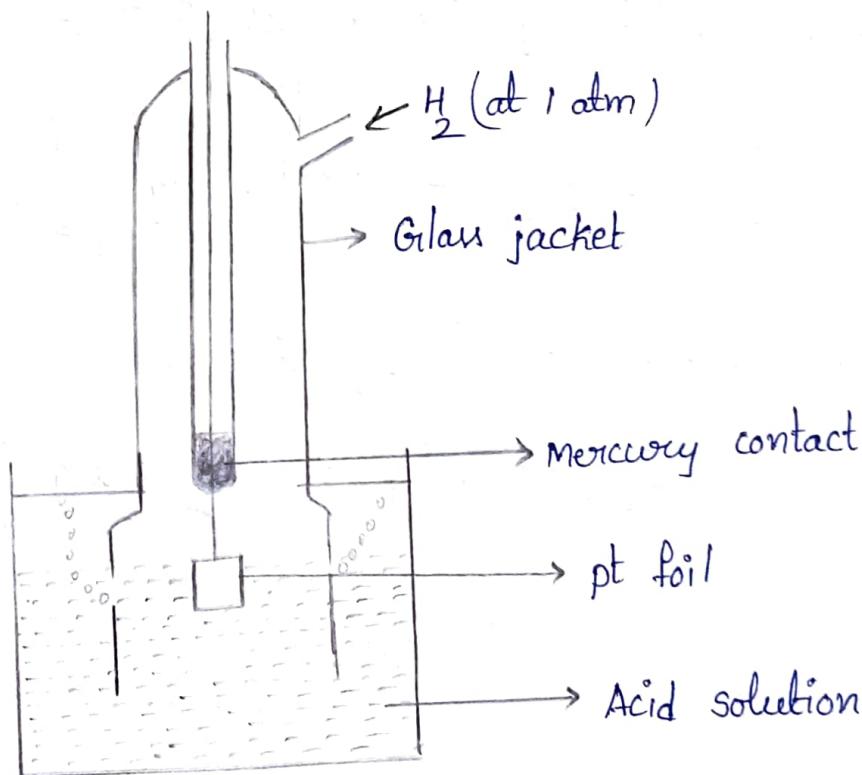
In general Nernst Equation is

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} + \frac{0.0591}{2} \log \frac{\text{Right hand side electrode}}{\text{Left hand side electrode}}$$

Reference Electrodes / Standard Electrodes :

- Reference Electrodes act as both anode and cathode.
- Reference electrode is the electrode the potential of which is known or arbitrary fixed at zero.
- The commonly used reference electrodes are
 - (1) H₂ (Hydrogen) Electrode
 - (2) calomel Electrode
 - (3) Ag / AgCl Electrode

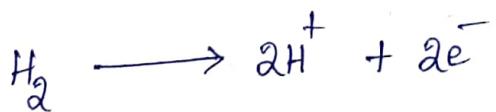
(a) Hydrogen electrode :



- It is the primary electrode. It consists of a small platinum strip coated with platinum black to absorb

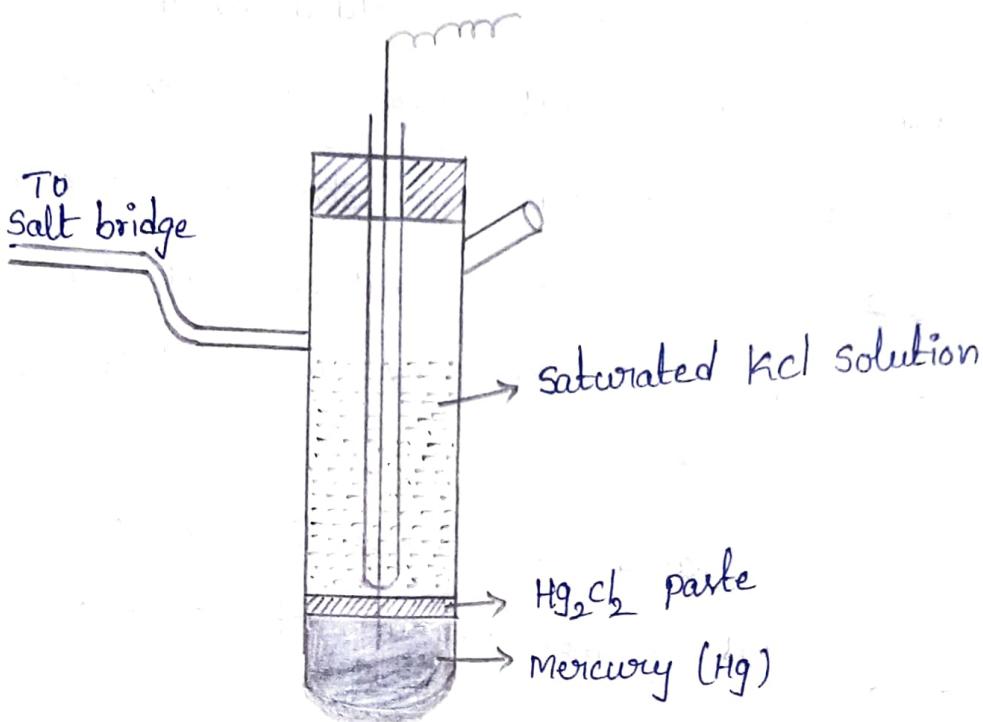
hydrogen gas. A platinum wire welded to the electrode and sealed through a glass tube makes contact with the outer circuit. The platinum strip is surrounded by an outer glass tube which has an inlet for hydrogen at the top and number of holes at the base for the escape of excess gas.

- The electrode is placed in a dilute solution of an acid. pure hydrogen gas is then passed into it at one atmosphere pressure.
- If the hydrogen act as anode,
A part of hydrogen gas absorbed by the platinised electrode, the excess is escapes thorough the lower holes.



- If the hydrogen electrode act as cathode,
- $$2H^+ + 2e^- \longrightarrow H_2 \uparrow$$
- The hydrogen electrode potential is fixed as Zero. sometimes Hydrogen electrode also gets 'poisoned' even if traces of impurities in the gas. To overcome this, some other electrodes called Secondary reference electrodes have been employed.

(2) calomel electrode : (Secondary Reference Electrode)

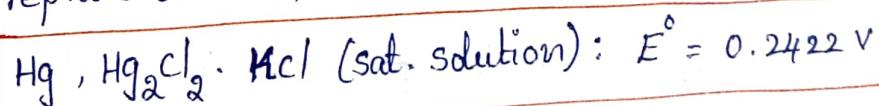


construction :

calomel electrode consists of a glass tube containing mercury at the bottom over which mercurous chloride is placed. The remaining portion of the tube is filled with a saturated solution of KCl.

The bottom of the tube is sealed with a platinum wire. The side tube is used for connecting with a salt bridge.

It is represented as,



- The single electrode potential of the calomel electrode is depending upon the concentration of KCl.

$$0.1 \text{ N KCl} = 0.3338 \text{ V}$$

$$1.0 \text{ N KCl} = 0.2800 \text{ V}$$

$$\text{saturated KCl} = 0.2422 \text{ V}$$

Anode reaction,

If the electrode acts as anode the reaction is



cathode reaction,

If the electrode acts as cathode the reaction is

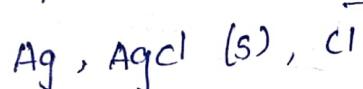


silver - silver chloride Electrode :

Secondary Reference Electrode :

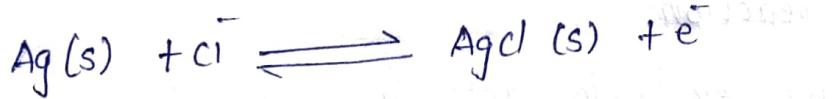
This electrode consists of a silver wire coated with a layer of silver chloride dipping into a solution of chloride (e.g. KCl or HCl) of unit activity.

This is represented as



- The standard electrode potential with respect to SHE is 0.2224 V at 298 K.

The electrode reaction is



The electrode potential is given by

$$E_{Ag/AgCl} = E_{AgCl}^{\circ} - \frac{RT}{nF} \ln \frac{1}{[Cl^-]}$$

- It is used to measure the pH of the solution.

Potential problems

$$- E_{cell} / EMF = E_R - E_{ox}$$

$$- E_{cell}^{\circ} / EMF = E_R^{\circ} - E_{ox}^{\circ}$$

$$- E_{ox} = E_{ox}^{\circ} + \frac{RT}{nF} \log \frac{[Red]}{[Ox]} \quad [Red=0]$$

$$- E_{Red} = E_{Red}^{\circ} + \frac{RT}{nF} \log \frac{[Red]}{[Ox]} \quad [Ox=0]$$

$$- E_{cell} = E_{cell}^{\circ} + \frac{RT}{nF} \log \frac{[Red]}{[Ox]}$$

$$- E_{cell} = E_{cell}^{\circ} + \frac{0.0591}{2} \log C$$

POTENTIOMETRIC TITRATIONS :

- potentiometric titration is very important application of emf measurement. In this method, a reference electrode is used.
- We know that the potential of an electrode dipping in a solution of an electrolyte depends upon the concentration of active ions (i.e. which changes the electrode potential)

$$E = E^{\circ} + \frac{RT}{nF} \log C$$

- A small change in the active ion concentration in the solution changes the electrode potential correspondingly. During the course of titration, the concentration of active ion decreases, thereby electrode potential of the indicator electrode decreases.
- thus, measurement of indicator electrode potential can serve as a good indicator of end-point or equivalence point of the titration reaction.
- potentiometric titrations are three types.
 - (i) Acid - Base potentiometric titration
(HCl - NaOH)

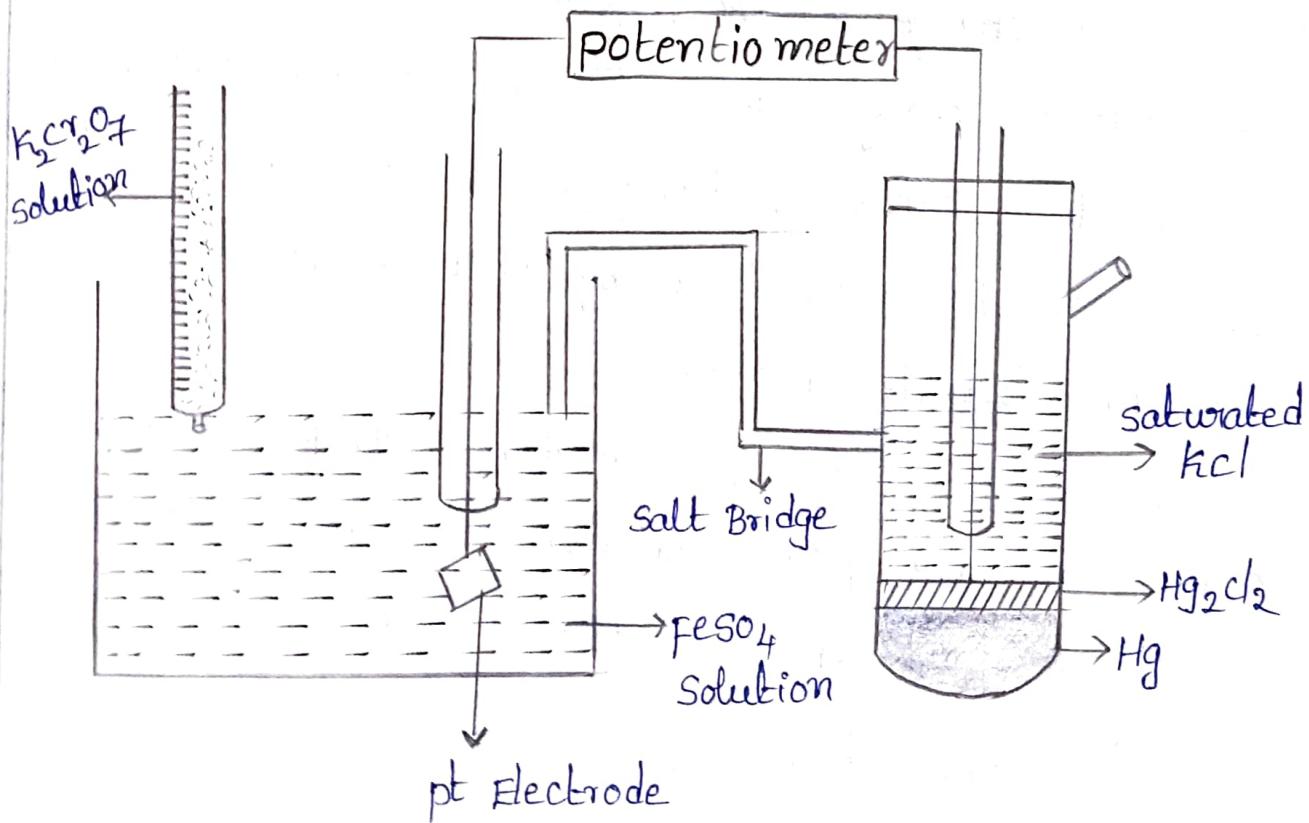
(2) Oxidation - Reduction (Redox)



(3) precipitation titrations ($\text{AgNO}_3 - \text{KCl}/\text{KI}$)

Oxidation - Reduction (Redox) :

- In this type of potentiometric titration, a reference electrode is saturated calomel electrode (SCE).
- Titrations involving oxidizing agent ($\text{K}_2\text{Cr}_2\text{O}_7$) and reducing agents can be followed potentiometrically by using platinum indicator electrode.



- known amount of FeSO_4 solution is taken in a beaker and the indicator electrode (platinum electrode) is inserted in it. It is then connected to a reference electrode (calomel electrode) to form a electro chemical cell. The cell is then connected to the potentiometer and its E_{cell} is determined.

- on adding $\text{K}_2\text{Cr}_2\text{O}_7$ from the burette, emf of the cell will increase first slowly, but at the equivalence point there will be sudden jump in potential, since change in the ratio of Fe^{+2} / Fe^{+3} ions concentration.

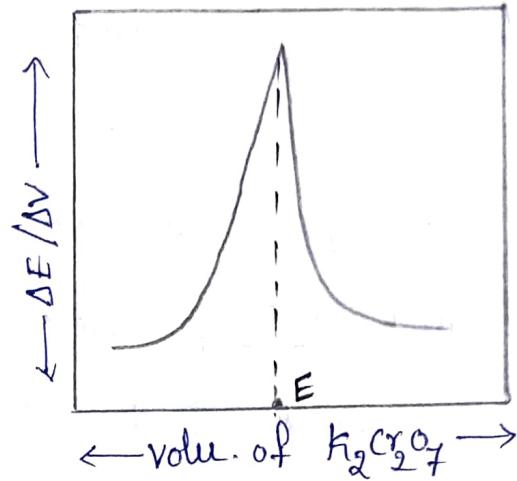
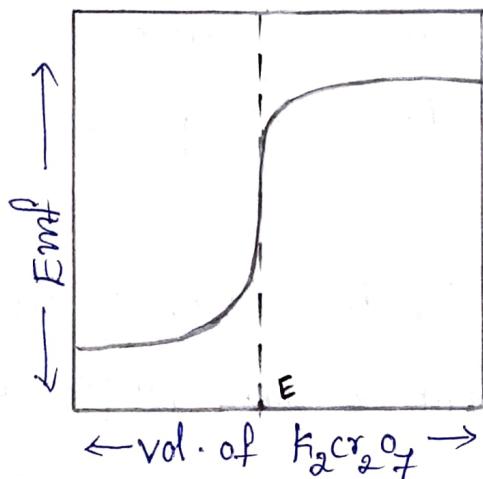


$$E = E^{\circ} + \frac{RT}{nF} \ln \frac{[\text{Fe}^{+3}]}{[\text{Fe}^{+2}]}$$

- When the emf is plotted against volume of $\text{K}_2\text{Cr}_2\text{O}_7$ added, a curve of the type shown in figure is obtained. The end point is the point, where the slope of the curve is maximum.

- A more sensitive and satisfactory method of detecting the end point will be the graph of $\Delta E/\Delta V$ against volume of $\text{K}_2\text{Cr}_2\text{O}_7$.

The resulting curve rises to a maximum at the equivalent point which is the end point.



Advantages of potentiometric titrations :

- The necessary apparatus required is cheap and easily available.
- This method can be used for coloured solution.
- Fixing up end point is easier when compared to the titrations in which indicators are used to fix up end point.
- very dilute solutions can be titrated with accuracy.
- Several components may be titrated in the same solution.

concept of conductivity :

Electrons play an important role when an electric current passes through a solid material. similarly ions (cations and anion) also play an important role when a current passes through a liquid such as an aqueous solution. since conductivity is an indicator of how easily current passes through a solution. The conductivity mainly depends on quantity of ions and their mobility in that solution.

conductance (C) :

The conductance of an electrolyte is the reciprocal of its resistance.

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

$$\text{units} = C = \frac{1}{\text{ohm}} = \text{ohm}^{-1} \text{ (or) mho.}$$

specific conductance (k) :

The reciprocal of specific resistance is called specific conductance. (or) It is the conductance of 1 cm^3 of a material (or) 1 volume of a liquid.

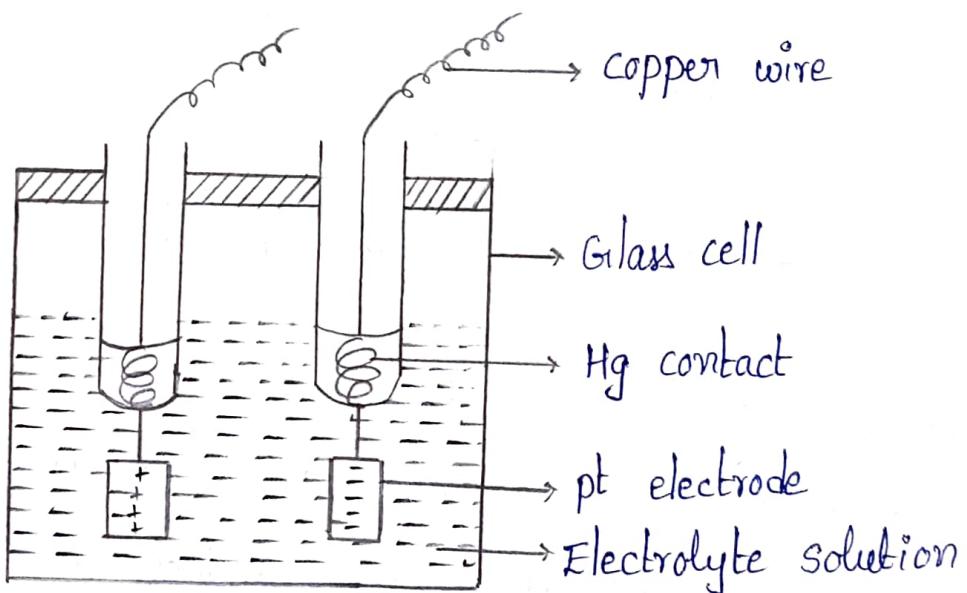
$$\text{units} = \text{ohm}^{-1} \text{ cm}^{-1} \text{ (or) mho cm}^{-1}.$$

conductivity cell :

The solution whose conductivity is to be determined is placed in a special type of cell "known as the conductivity cell". Thus conductivity cell is a cell used to measure the conductivity of the solution.

construction :

A conductivity cell consists two electrodes, fitted in the cell, made of platinum plates coated with platinum black. These are welded to platinum wires fused in two thin glass tubes and are connected to copper wires.



conductivity cell

Conductometric Titration :

principle : conductometric titration is a volumetric method based on the measurement of conductance of the solution during the titration.

The conductance of a solution depends on

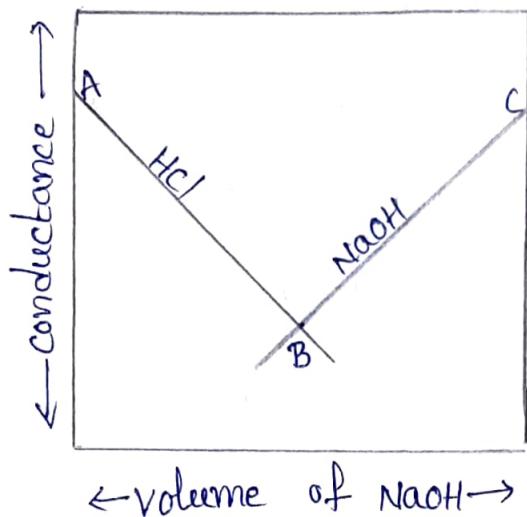
- ii) The number of free ions
- iii) The mobility of the ions.

Acid - Base Titrations :

(ii) strong acid vs strong base (HCl vs NaOH)

known amount of acid (HCl) is taken in the beaker and conductivity cell is immersed and the alkali (NaOH) in the burette. Initially the conductivity of the HCl is high, this is due to the presence of fast moving H^+ ions (point A in the graph). As the NaOH is added gradually, conductance will be going on decreasing until the acid has been completely neutralized (indicated by the line AB). This is due to decrease of the H^+ ions in the solution. The point 'B' indicates complete neutralization of all H^+ ions.





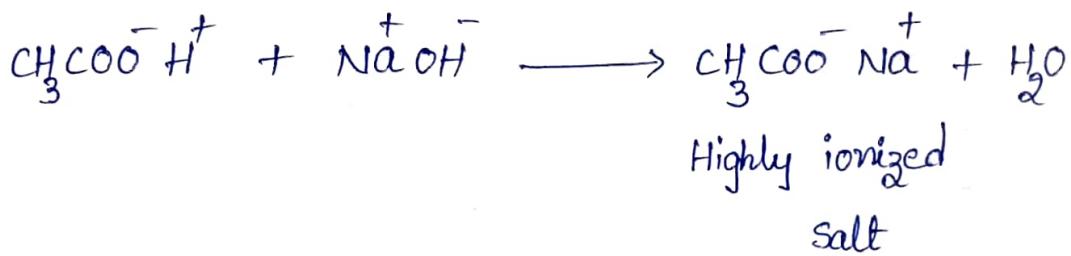
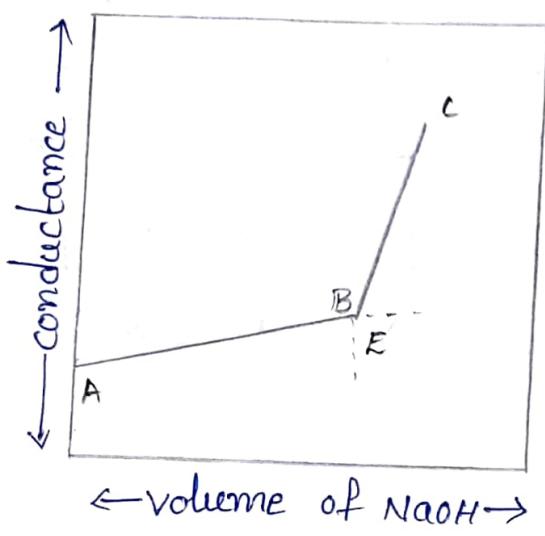
further addition of NaOH will introduce the fast moving ions. Therefore the conductance, after reaching a certain minimum value, will begin to increase (indicated by the line BC).

on plotting the conductance against the volume of alkali added, the two lines intersect at a point 'B' gives the end point. This corresponds to the volume of NaOH required for neutralization.

a. Weak acid vs Strong Base (CH_3COOH vs NaOH)

known amount of weak acid (CH_3COOH) is taken in the conductivity cell and the alkali (NaOH) in the burette. Initially the conductivity of CH_3COOH is low, this is due to the poor dissociation of CH_3COOH . (point A in the graph). As NaOH is added gradually, conductance will be going on slowly increasing

(indicated by the line AB). This is due to the formation of highly ionised $\text{CH}_3\text{COO}^-\text{Na}^+$.



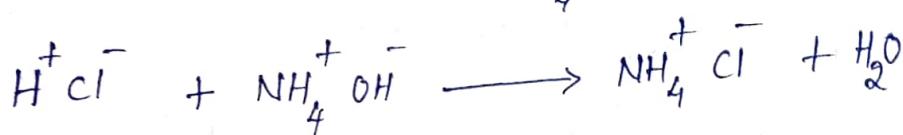
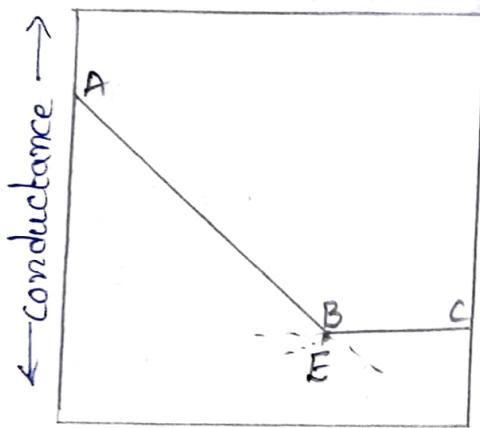
The point 'B' indicates complete neutralization of CH_3COOH .

Further addition of NaOH introduces excess of fast moving OH^- ions. Therefore the conductance of the solution begins to increase even more sharply than before (line AB).

On plotting the conductance against the volume of alkali added, the two lines intersect at point B gives the end point.

(3) strong acid vs weak base (HCl vs NH₄OH)

known amount of HCl is taken in the conductivity cell and NH₄OH is added gradually from the burette. Initially the conductivity of HCl is high (point A in the graph). As the NH₄OH is added gradually the conductance will be going on decreasing until all the acids been completely neutralised (line AB). This is due to the decrease of the H⁺ ions in the solution.

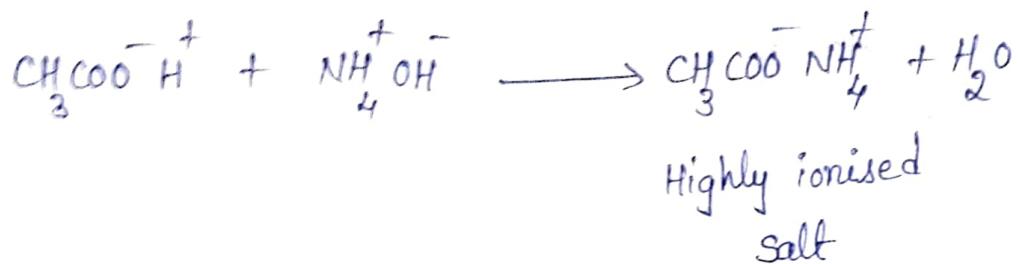


further addition of poorly ionized NH₄OH does not cause any appreciable change in the conductance (line BC). The point of intersection 'B' is the end point.

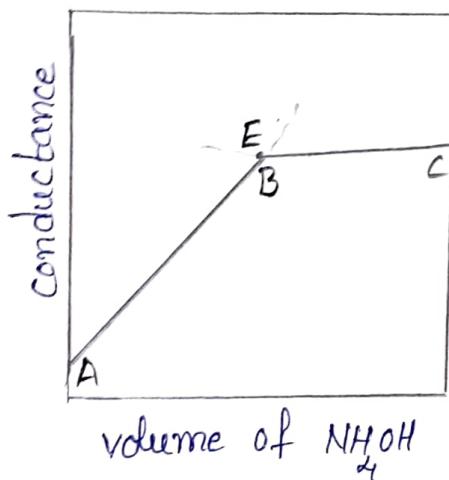
4. Weak acid vs Weak Base (CH₃COOH vs NH₄OH)

When CH₃COOH is titrated against NH₄OH, conductivity is going on increasing (indicated by

by the line AB). This is due to the formation of highly ionised salt $\text{CH}_3\text{COO}^- \text{NH}_4^+$.



After the neutralisation of CH_3COOH , further addition of poorly ionized NH_4OH does not cause any appreciable change in the conductance. (line BC). The point of intersection 'B' is the end point.



Electro chemical sensors :

Sensor is the device which gives information about the sample. (or)

Sensor are devices which convert one form of energy into another.

Electro chemical Sensors are two types.

(1) potentiometric Sensor

(2) Amperometric Sensor

(1) potentiometric sensor : potentiometry deals with the electromotive force (emf) generated in an galvanic cell, where spontaneous chemical reaction is taking place.

Analytical potentiometry deals with the use of the emf response of a potentiometric cell to determine concentrations of analytes in samples.

Ex:- clark oxygen Sensor.

(2) Amperometric Sensor :

Amperometric sensors are gives information from the current - concentration relationship alone.

- Amperometric measurements are made by recording a current flow in a cell at a single applied potential.

Ex - Glucose Sensor.

Oxygen Sensors

- the dissolved oxygen in a variety of aqueous environments, such as seawater, blood, sewage and soils can be determined by using clark oxygen sensor.

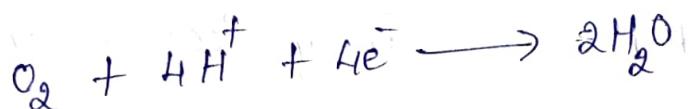
clark oxygen sensor:-

- The cell consists of a platinum disk cathodic working electrode imbedded in a centrally located cylindrical insulator, surrounding the lower end of this insulator is a ring - shaped silver anode.
The tubular insulator and electrodes are inserted in a second cylinder that contains a buffered solution of potassium chloride (KCl).
- A thin ($\sim 20 \mu\text{m}$) replaceable, oxygen permeable membrane of teflon (or) polyethylene is held in place at the bottom end of the tube by an O-ring. The thickness of the electrolyte solution between the cathode and the membrane is approximately $\sim 10 \mu\text{m}$.
- When the oxygen sensor is immersed in a flowing solution of the analyte, oxygen diffuses through the membrane into the thin layer of electrolyte

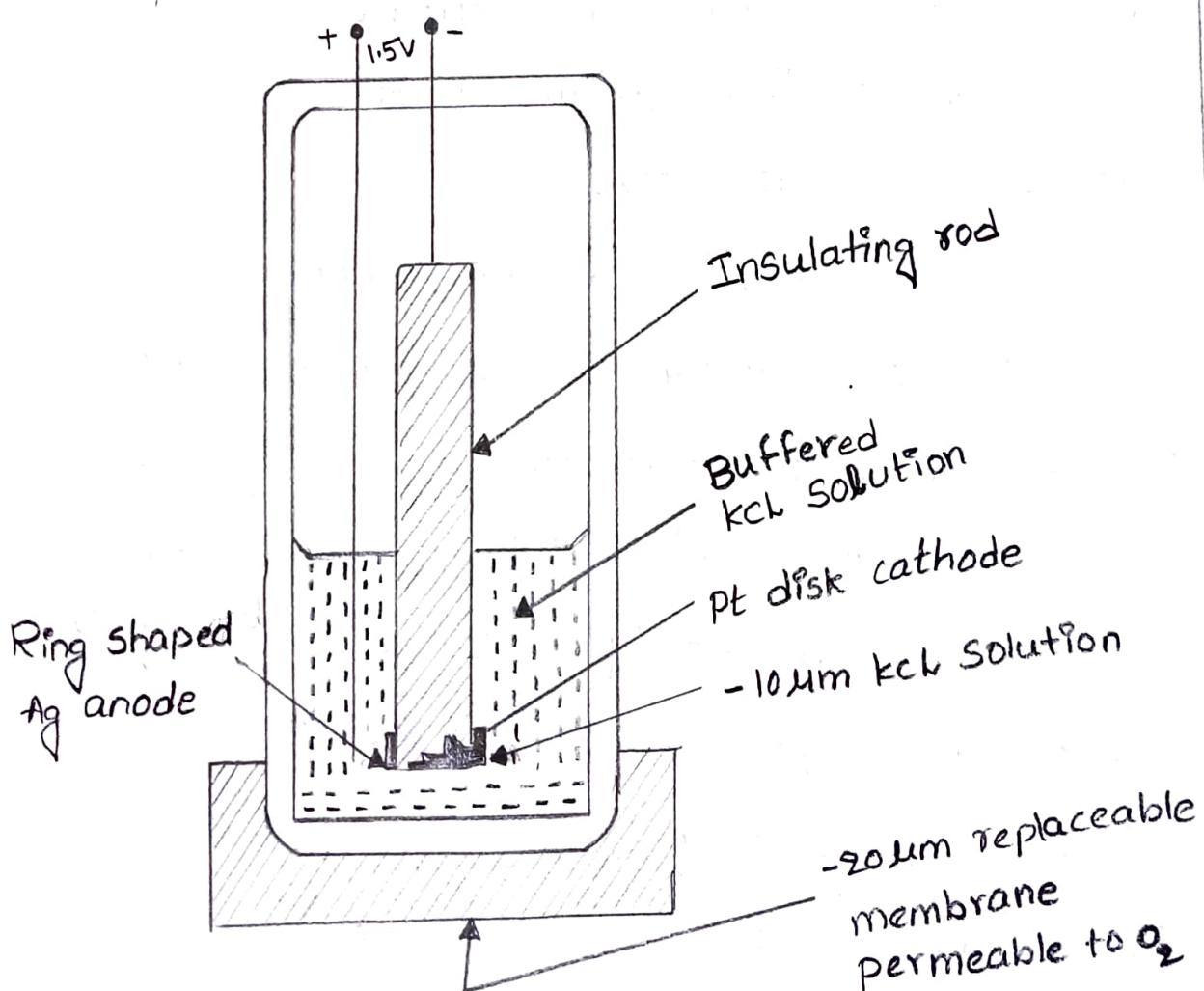
immediately, where it diffuses to the electrode and is immediately reduced to water. This reaction takes place in a reasonable period 10 to 20s.

under these conditions, it is the rate of equilibration of the transfer of oxygen across the membrane that determines the rate at which currents are achieved.

cathodic reaction



Anodic reaction



clark Electrochemical oxygen sensor

Enzyme - Based Sensors

Glucose Sensor

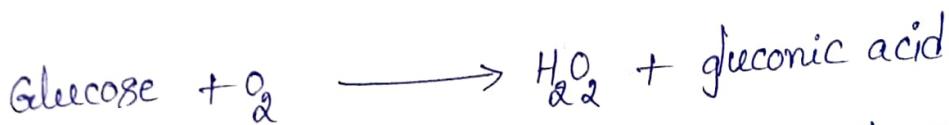
- the glucose sensor is widely used in clinical laboratories for the determination of glucose in blood. This device is similar in construction to oxygen sensor.

- The membrane in this case is more complex and consists of three layer. The outer layer is a polycarbonate film that is permeable to glucose but impermeable to proteins and other constituents of blood.

- the middle layer is an immobilized enzyme, glucose oxidase in this instance.

- The inner layer is a cellulose acetate membrane, which is permeable to small molecules, such as hydrogen peroxide.

- When this device is immersed in a glucose-containing solution, glucose diffuses through the outer membrane into the immobilized enzyme, where the following catalytic reaction occurs.



- the hydrogen peroxide then diffuses through the inner layer of membrane and to the electrode

surface, where it is oxidised to give oxygen.



Battery (or) cells

A battery is an arrangement of several electrochemical cells connected in a series, that can be used as a source of direct electric current.

Thus

A cell contains only one anode and cathode.

A battery contains several anodes and cathodes.

Types of Battery :

(1) primary Battery (or) primary cells :

In these cells, the electrode and the electrode reactions cannot be reversed by passing an external electrical energy. The reactions occur only once and after use they become dead. Therefore they are not chargeable.

Ex : Zinc-air battery, Dry cell, Li primary

battery.

(2) Secondary Battery (or) Secondary cells :

In these cells, the electrode reactions can be reversed by passing an external electrical energy.

therefore, they can be recharged by passing electric current and used again and again. These are also called storage cells.

Ex:

3) Flow battery (or) fuel cells :

In these cells, the reactants, products and electrolytes are continuously passing through the cell. In this chemical energy gets converted into electrical energy.

Ex: $H_2 - O_2$ fuel cell, Methanol fuel cell.

Primary Battery :

Zinc - Air battery :

Zinc - air batteries work by the oxidation of Zn with oxygen from the air.

construction :

A mass of loose granulated zinc particles mixed with an electrolyte such as KOH to form a porous cathode. The Zinc plates, in the other side, act as anode.

working :

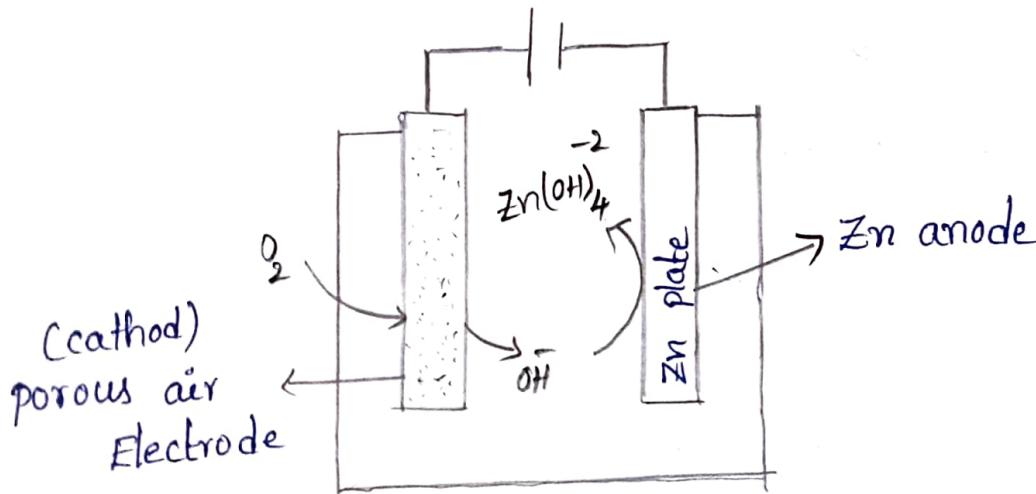
Oxygen from the air, at the cathode forms hydroxide ions which migrate into the Zinc paste and form zincate i.e $[Zn(OH)_4]^{2-}$ releasing electrons

to travel to the cathode. Thus Zincate decays into Zinc oxide and water.

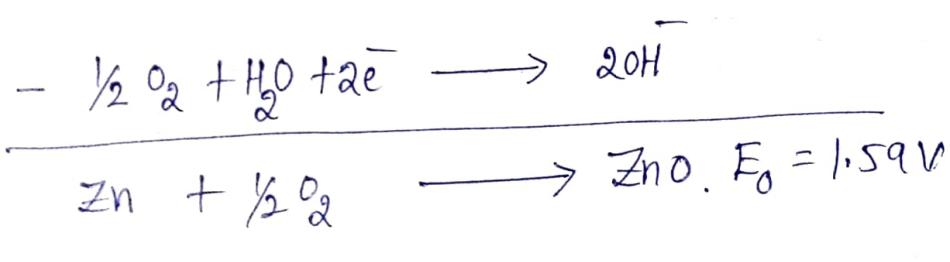
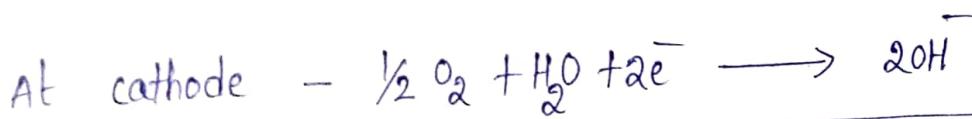
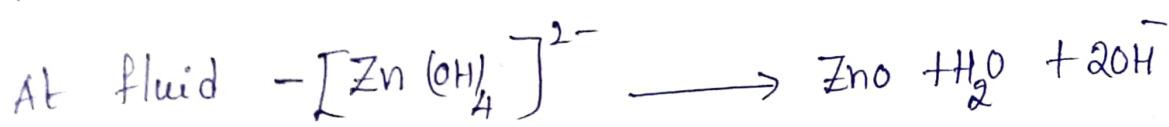
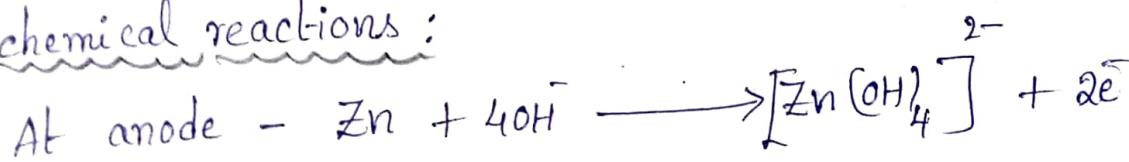
The water and hydroxide ions from the electrodes are recycled, so the water is not consumed. This battery produces 1.65 V.

Factors affecting rate of reaction

1. The rate of reaction can be controlled by varying the oxygen flow.
2. The rate of reaction can be increased by replacing oxidized zinc/electrolyte paste by fresh paste.



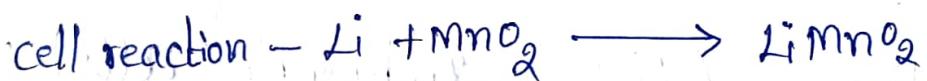
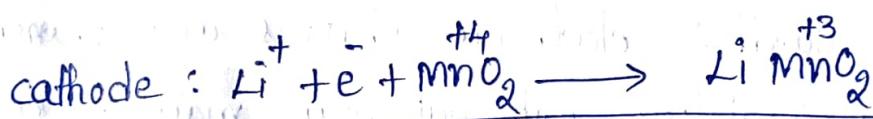
Chemical reactions:



(2) Li primary cells:

(17)

- The most common Li primary cell is the lithium manganese dioxide battery (LiMnO_2)
- Anode is Li and cathode is MnO_2 which is heated above 300°C to remove H_2O before incorporated in the cathode, thus increases the efficiency of cell.

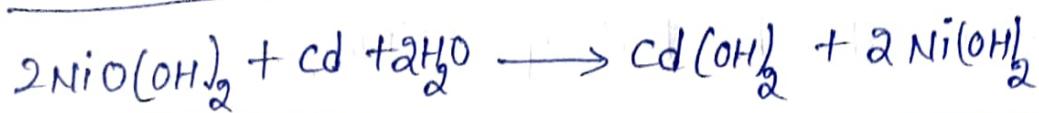
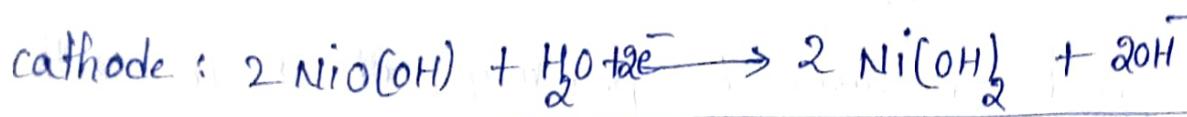


- this cell produces a voltage of about 3.4V which is more than twice that of dry cell due to light weigh of Li.
- they are widely used in applications that require higher current drain like fully automatic cameras, electronic devices such as calculators and watches.

Secondary batteries:

ni-cd battery / Nicad battery:

This is recently developed, portable, rechargeable cell and its voltage is fairly constant (about 1.4V). It consists of Cd anode, a cathode composed of a paste of $\text{NiO(OH)}_2(s)$.



- this reaction can be readily reversed, because the reaction products, $\text{Cd}(\text{OH})_2$ and Ni(OH)_2 , adhere the electrode surfaces.

- They are mainly used in electronic calculators, CD players, electronic cars, cordless electronic shavers and transistors.

- they have long durability with out any maintenance.

(2) Li Ion battery : (Li Ion Rechargeable cells) :

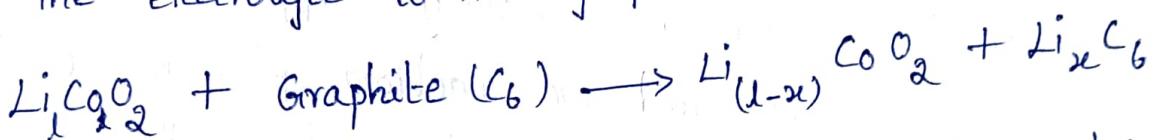
- These cells do not contain metallic 'Li' instead it contains lithium ions (Li^+)

Anode : carbon compound, graphite

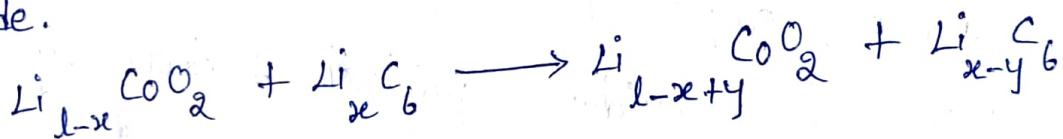
cathode : Lithium cobalt oxide

- In these cells the true oxidation / reduction does not takes place, instead of that transport of Li^+ ions through the electrolyte from one electrode to the other takes place.

- When the cell is constructed, it is in its uncharged state with no Li^+ ions between the layers of 'c' atoms in the graphite.
- When the cell is charging, Li^+ ions leave LiCoO_2 and travel through the electrolyte to the graphite C_6 .

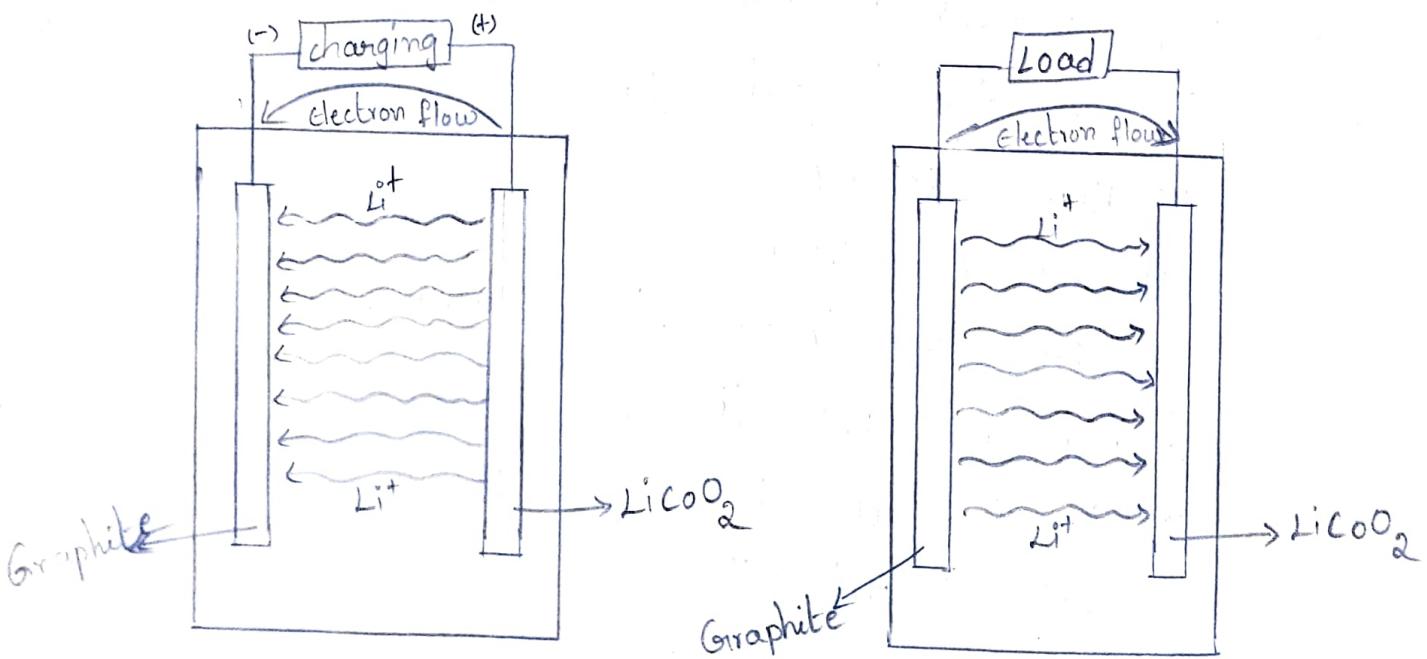


When the cell spontaneously discharges to provide electrical power Li^+ ions move back through the electrolyte to the cobalt oxide while electrons move through the external circuit from the graphite electrode to the cobalt oxide electrode.



y = amount of Li^+ transferring from Li_xC_6 to LiCoO_2

x = Transferring amount of Li^+ from LiCoO_2 to graphite.



- Thus the charging and discharging cycle simply sweep lithium ions back and forth between two electrodes with electrons flowing through external circuit to keep the charge balance.

Fuel cells / Flow battery :

The reactants are continuously supplied from an external source to the cell and the reaction products continuously removed, such cells are called as fuel cells.

- In a fuel cell electrical energy is obtained without combustion from O_2 and gas can be oxidized. Hence fuel cell converts chemical energy of the fuel directly to electricity.

The important steps in the fuel cell



In which fuel is gas / liquid and oxygen being oxidized. In fuel cell the chemical energy is provided by fuel and easily converts to electrical energy.

- Fuel cells are characterized by,

(1) High efficiency

(2) Low noise levels

(3) free from vibration, heat transfer.

- The most important fuel cells are

(1) $H_2 - O_2$ fuel cell

(2) methanol fuel cell

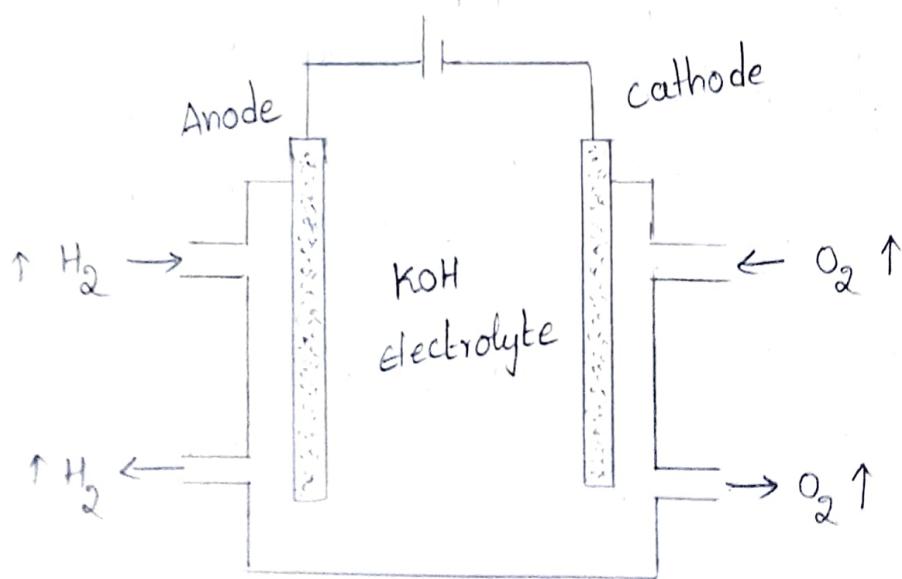
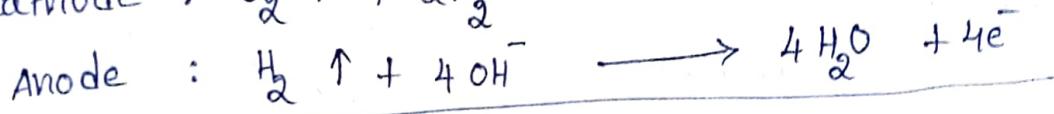
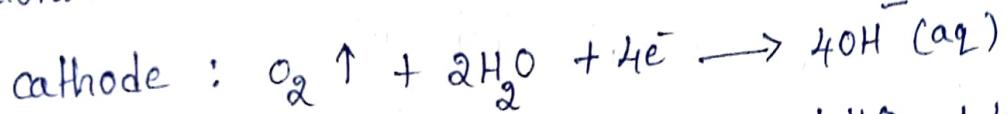
(3) solid oxide fuel cell

$H_2 - O_2$ fuel cell / Alkaline Fuel cell :

- It is successful and simplest cell.

electrodes : porous platinum electrode, electrolyte 35-40% KOH.

- It consists of hot concentration KOH electrolyte solution and 2 porous pt electrodes are dipped in KOH solution.
- At anode H_2 gas and at cathode O_2 gas is supplied, which are come in contact with electrodes and shows following reactions.



Application :

- iii) They are used as energy source in space vehicle, submarines or military vehicles.

- (2) The reaction product is water, which is harmless and used as source of fresh water to the astronauts.
- (3) The weight of fuel battery for 15 days is about 250 kg.

Disadvantages :

- (1) Energy cost of generating H₂ fuel.
- (2) Difficult in storage and distribution of highly flammable.
- (3) Availability of alternative fuels.

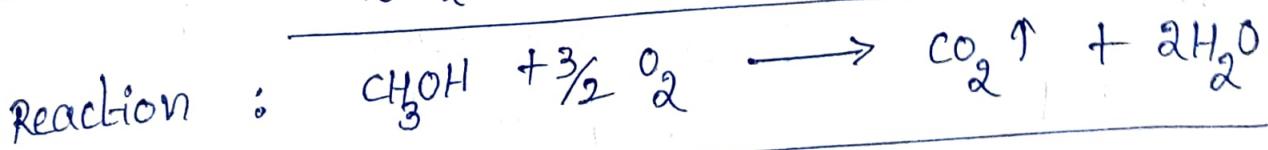
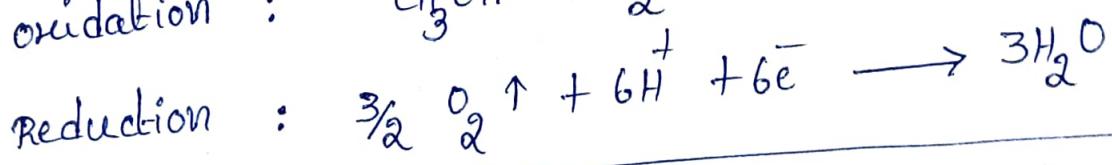
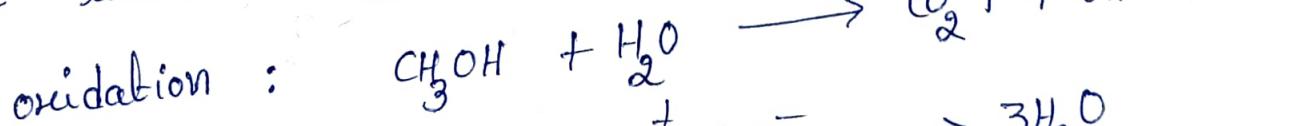
(2) Methanol fuel cell / proton exchange fuel cell :

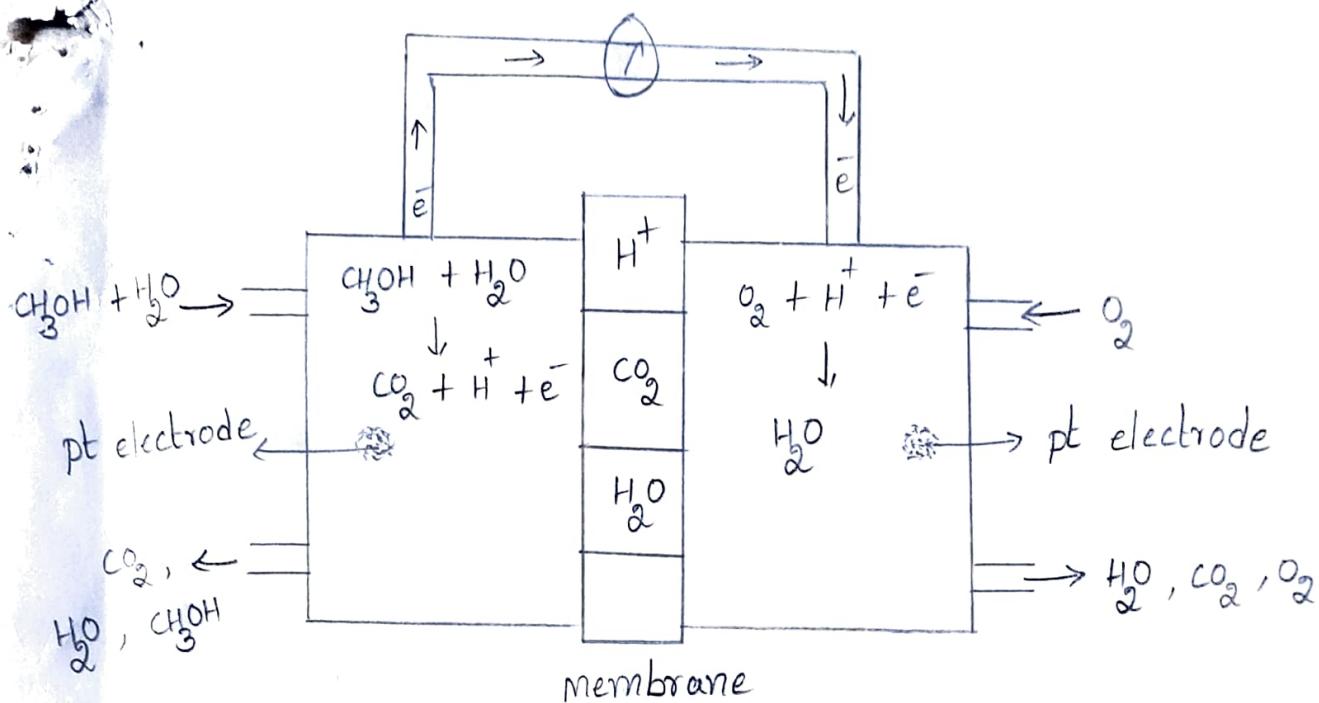
Electrodes : porous pt electrodes, separated by proton transport membrane.

Anode : oxidation of CH₃OH and H₂O absorbed on catalyst to form CO₂ with loss of H⁺ (protons) and e⁻s.

Cathode : H₂O is liberated by reduction of O₂, which accepts electrons and protons. Protons are transported from anode to cathode through proton transport membrane, while electrons transport through electric circuit.

The liberated H₂O (at cathode) is consumed at anode.





Advantages :

- (1) storage of methanol is much easier than H_2 (does not require high pressure or low temperature) since methanol is liquid form -97.0°C to 64.7°C (-142.6°F to 148.5°F)
- (2) The energy density of methanol is greater than H_2 .

Disadvantages :

- (1) During methanol oxidation, CO formed which is strongly adhere on to pt catalyst, reducing the surface area and lowering the performance.
- (2) CH_3OH is toxic and flammable and so most care is necessary.
- (3) Limited power is produced by them.