

UNIT-III

Electrochemistry and Applications

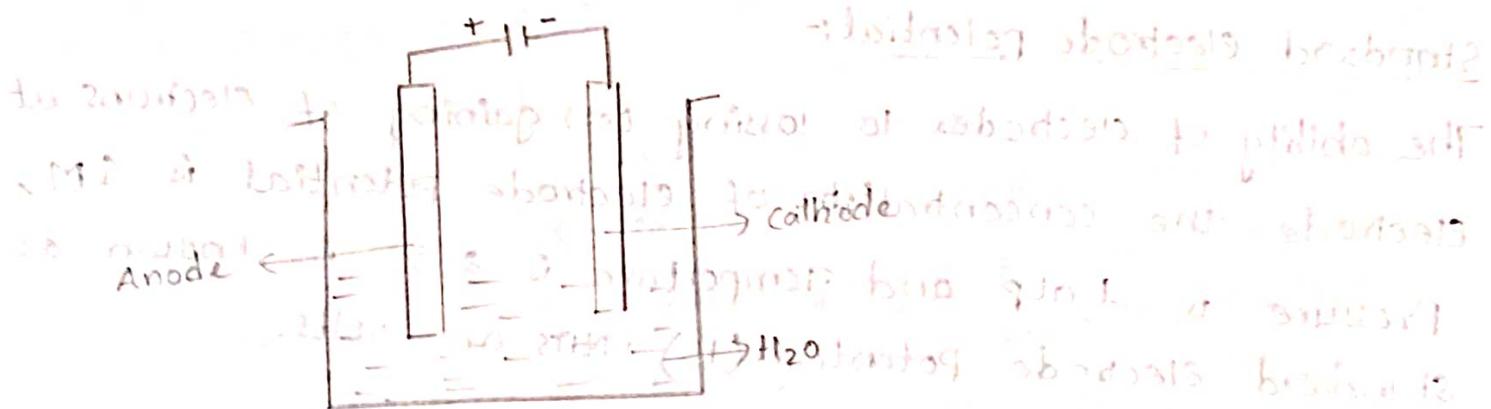
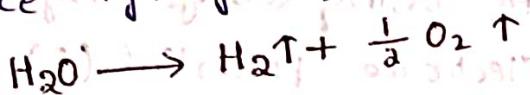
It is a branch of physical chemistry, which deals the relationship between electricity and chemical effects.

In 1800, William Nicholson observed electricity is related to chemical effects in electrolysis of water.

Activity - 1

electrolysis of water :-

When ever we supply electrical energy into water sample which produce hydrogen gas and oxygen gas.

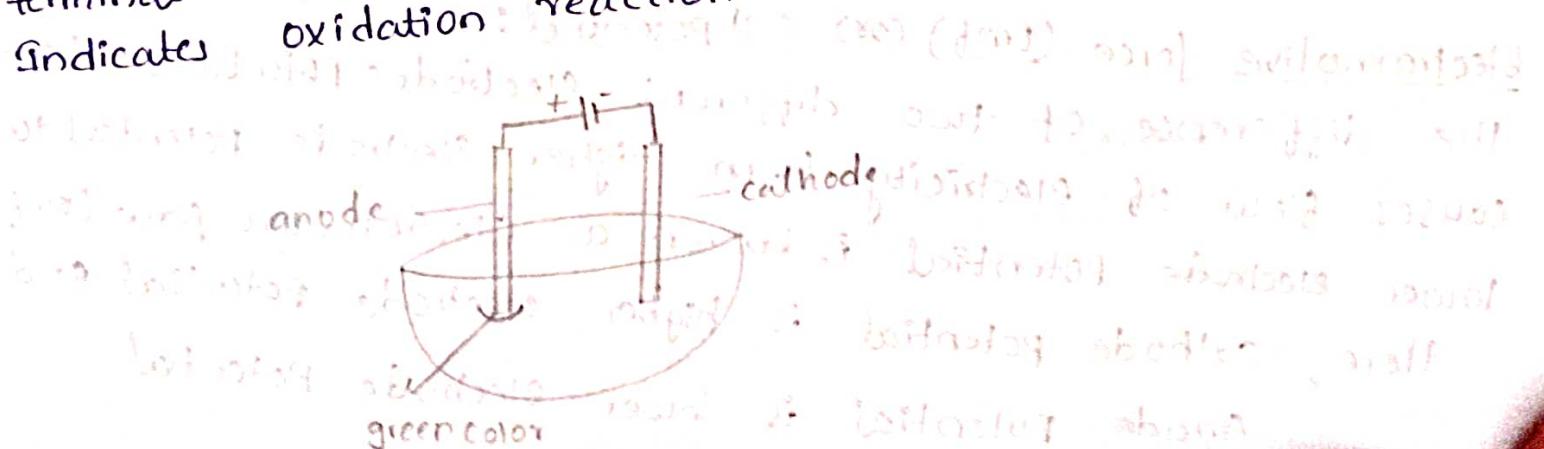


Activity - 2

Potato activity :-

In this activity we will take half portion big size potato, two graphite rods into potato further both are connected

Placing two graphite rods after some time the positive terminal dipped position converted into green colour which indicates oxidation reaction at Anode portion.



Oxidation :-

The addition of oxygens (or) Removal of Hydrogen (or)
Removal of electrons takes places in a chemical reaction is known as oxidation.

Reduction :-

The Removal of oxygens (or) addition of Hydrogen (or)
Removal of electrons takes places in a chemical reaction is known as reduction.

Electrode potential :-

The ability of electrodes to lossing (or) gaining of electrons at electrodes is known as electrode potential. (E) units are volts.

Standard electrode potential :-

The ability of electrodes to lossing (or) gaining of electrons at electrode the concentration of electrode potential is 1M, Pressure is 1 atm and Temperature is 25°C is known as standard electrode potential. (E_0) units are volts.

Oxidation potential :-

The ability of electrodes to lossing of electrons at electrode is known as oxidation potential. (E_{Oxy})

Reduction potential :-

The ability of electrodes to gaining of electrons at electrode is known as Reduction potential (E_{Red})

Electromotive force (EMF) (or) cell potential :-

The difference of two different electrodes potentials which causes flow of electricity from higher electrode Potential to lower electrode Potential is known as Electromotive force (emf)
Here, cathode potential is higher electrode potential and Anode potential is lower electrode potential

$$\text{EMF (or) } E_{\text{cell}} = E_{\text{cathode}} - E_{\text{Anode}}$$

Reference electrode:-

The electrode which consisting of known potential value is known as reference electrodes.

Base upon potential values the Reference electrodes are two types

Those are. 1. Primary Reference electrode

2. Secondary Reference electrode.

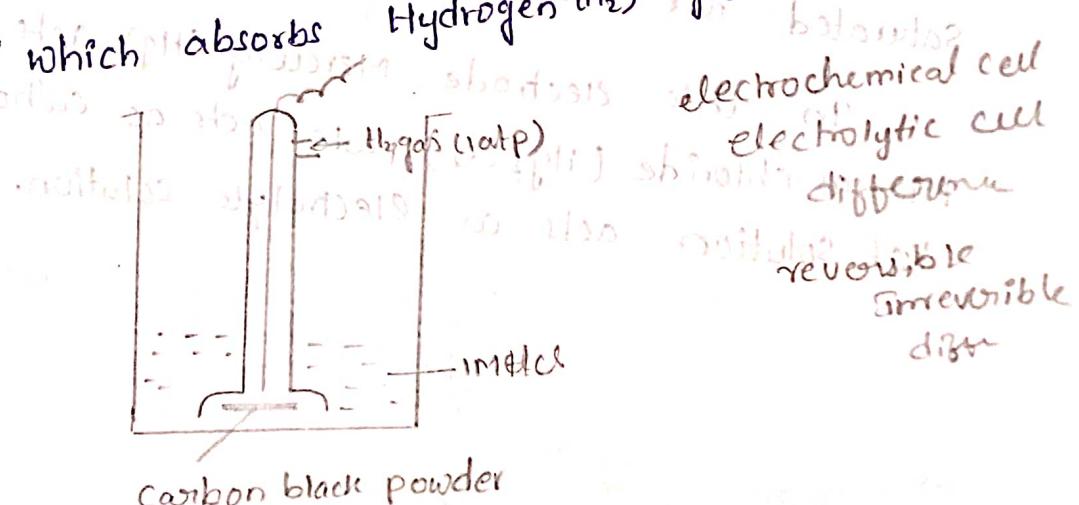
1. Primary Reference electrodes:-

The electrodes which consisting of zero potential value is known as Primary Reference electrodes.

Example:- Normal Hydrogen electrode (or) standard Hydrogen electrode (NHE (or) SHE)

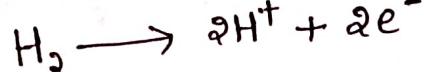
Normal Hydrogen electrode (NHE)

In a construction of electrode placing a platinum foil into 1M HCl solution covered with open bell shaped glass tube consisting of one inlet chamber. through the inlet chamber we will supply Hydrogen gas at 1atp into electrolyte solution. The platinum foil is coated with 'carbon black powder' which absorbs Hydrogen (H_2) gas



At Anode:-

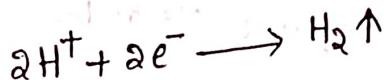
Hydrogen gas losses two electrons to produce H^+ ions.
Here, lossing of electrons indicates oxidation reaction



At Cathode :-

According to Arhenius Acid base theory HCl as strong acid which produce more no. of H^+ ions which gains two electrons to form Hydrogen gas. Two electrons to form Hydrogen gas indicates Reduction Reaction.

Here, gaining of electrons indicates Reduction Reaction.



Secondary Reference Electrode :-

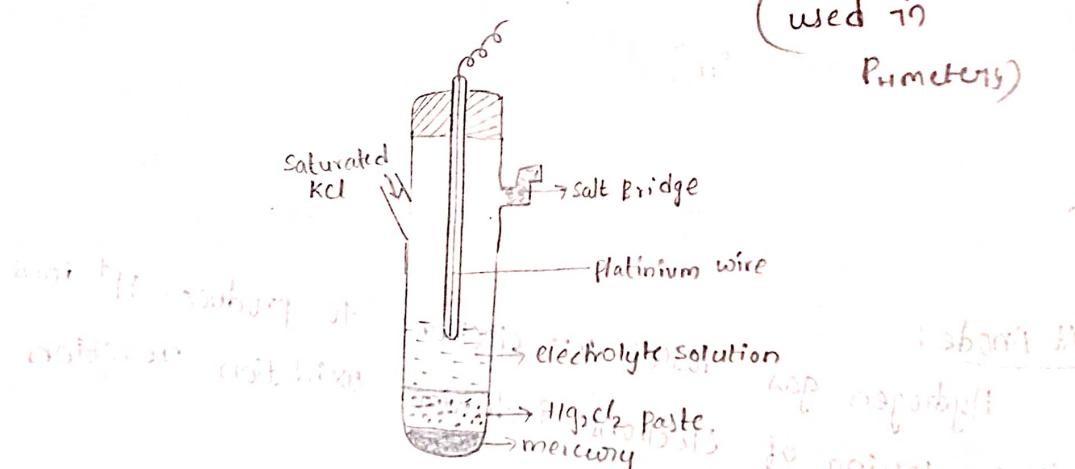
The electrode which consisting of constant (or) fixed potential values based upon electrolyte solution is known as Secondary Reference electrodes.

Example :-

1. Calamal electrode (or) Hg/Hg_2Cl_2 electrode

In the construction of Calamal electrode we will take cylindrical glass tube having three inlet chamber. In one inlet chamber we will hang platinum wire, The another inlet chambers we will add salt bridge and saturated KCl solution. In this electrode Mercury (Hg) acts as Anode and Merscurys chloride (Hg_2Cl_2) paste acts as cathode and saturated KCl solution acts as electrolyte solution.

(used in
potentiometry)

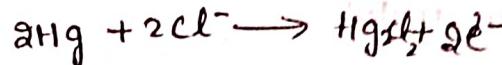


At Anode :-

Mercury acts as anode which are treated with chloride ions from the electrolyte solution which losses two electrons.

to form Mercurous chloride (Hg_2Cl_2) paste.

Here, lossing of electrons \rightarrow Indicates oxidation reaction.



At cathode :-

Hg_2Cl_2 acts as cathode which gains two electrons to from mercury metal and chlorine ions.

Here, gaining of electrons indicates reduction reaction.



The electrode potential of calomel electrode will depends upon nature of the electrolyte solution.

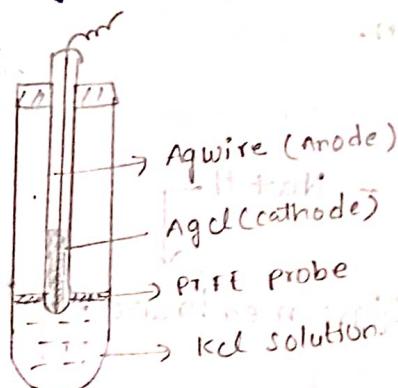
$$0.1 \text{ M KCl} \rightarrow 0.375 \text{ V}$$

$$1 \text{ M KCl} \rightarrow 0.280 \text{ V}$$

$$\text{saturated KCl} \rightarrow 0.272 \text{ V}$$

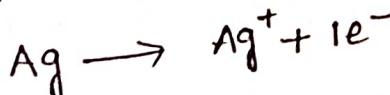
Silver | silver chloride ($Ag/AgCl$) electrode:

In the construction of silver-silver chloride electrode 'Ag' metal acts as anode, 'AgCl' acts as cathode and KCl solution acts as electrolyte, Poly Tetra Fluoro ethene (PTFE) probe acts as a separator.



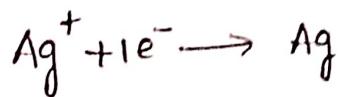
At anode:-

'Ag' metal acts as anode which loses electrons to from Ag^{+} ion. Here, lossing of electrons indicates oxidation reaction.



At cathode:-

"AgCl" dissociated into "Ag" ions which are gains electrons to form (Ag)silver metal.



Here,

The electrode potential of Ag-AgCl electrode in glass membrane is 0.228V

This electrode acts as inert electrode in glass membrane.

Glass membrane electrode / ion selective electrode:-

In the construction of glass membrane electrode, Ag-AgCl electrode is inert

Platinum wire is indicator

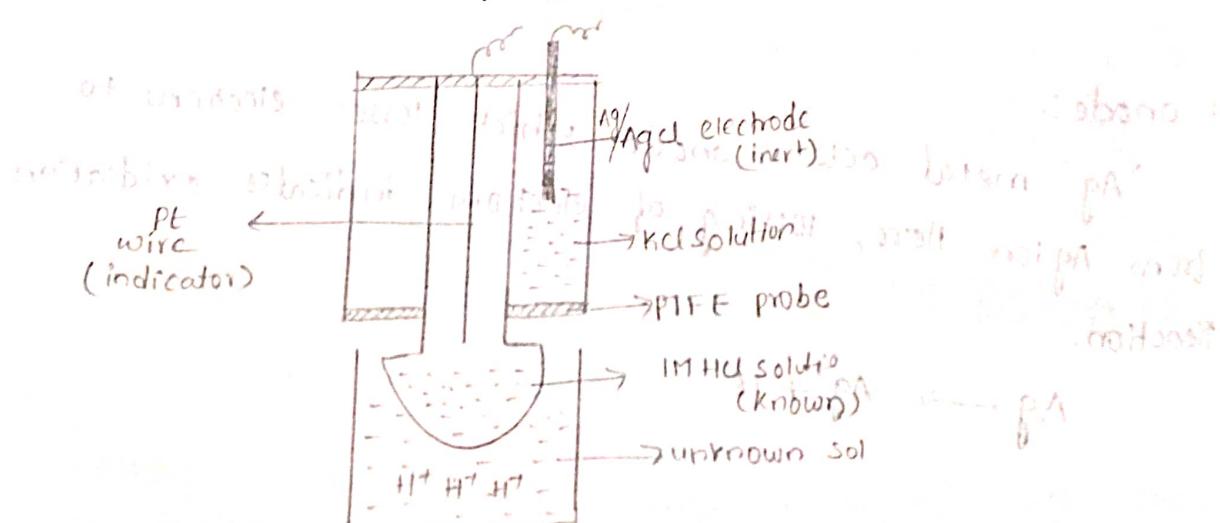
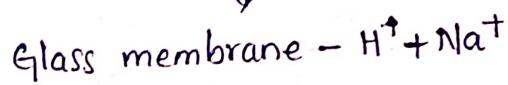
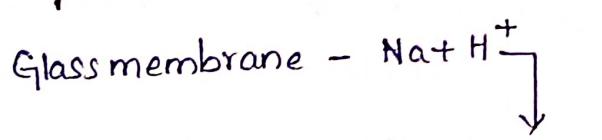
and glass membrane is ion exchanger

The composition of glass membrane electrode is

72% SiO_2 , 22% Na_2O and 6% of CaO .

Inside the glass membrane we will take one 'M' HCl (1M HCl) as a known solution and container we should take unknown solution.

ion-exchange reaction.



By using glass membrane electrode we calculate pH of the unknown solution which shows acidic, basic or neutral nature.

$$E_{\text{q}} = E_{\text{q}}^{\circ} + 0.0591 \text{ pH}$$

$$\text{pH} = \frac{E_{\text{q}} - E_{\text{q}}^{\circ}}{0.0591}$$

where, E_{q} = electrode potential of glass membrane

E_{q}° = standard electrode potential of glass membrane

~~nm~~ Electro chemical cell:-

A device which converts chemical energy into electrical energy and electrical energy into chemical energy by spontaneous redox reaction is known as electrochemical cell.

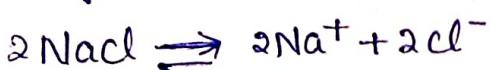
→ On based upon energy conversion electrochemical cells are divided into two types.

1. electrolytic cell

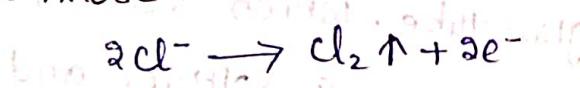
2. Galvanic cell

Electrolytic cell:-

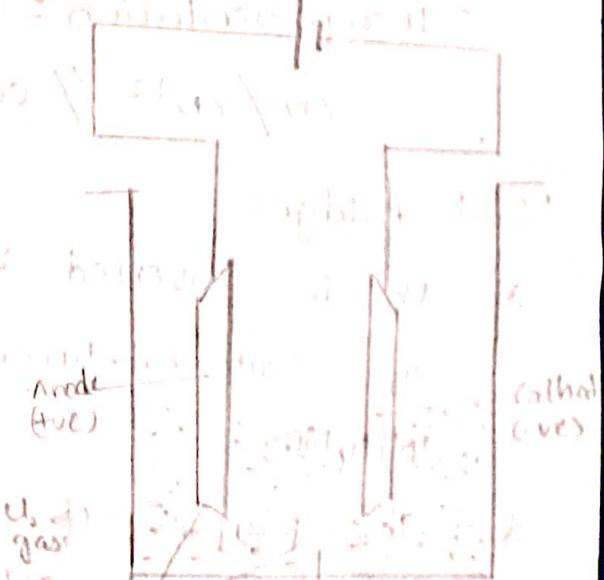
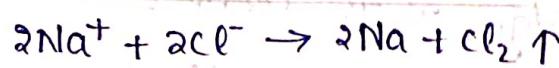
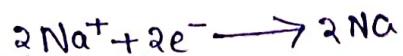
A device which converts electrical energy into chemical energy by spontaneous redox reaction is known as electrolytic cell.



At Anode:-

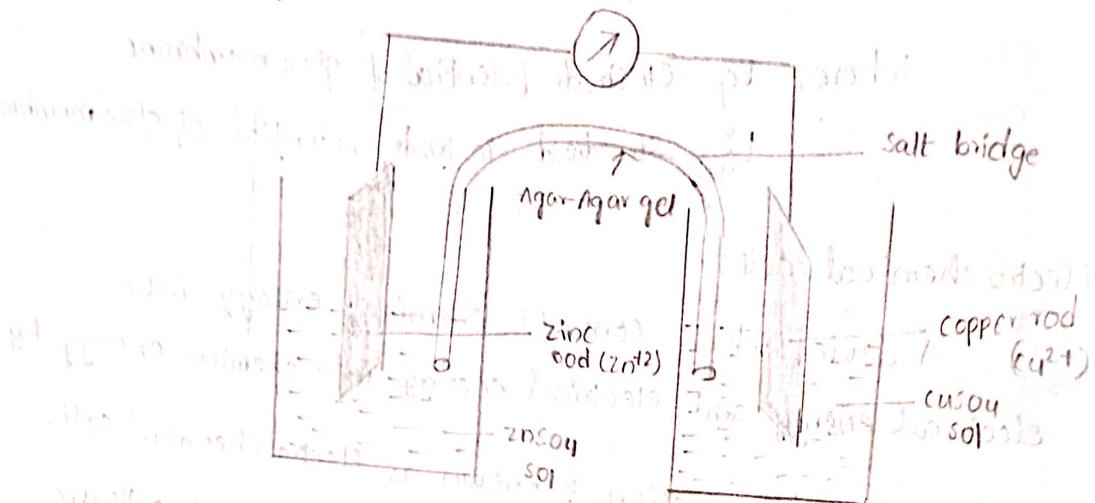


At cathode:-

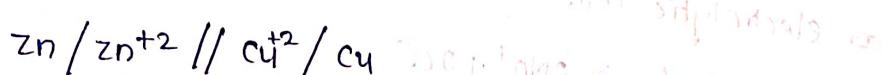


Galvanic cell (or) voltaic cell (or) Daniel cell

The device which converts chemical energy into electrical energy by spontaneous redox reaction is known as galvanic cell.



Cell representation :- (cell notation)



Salt bridge:-

- * It is Inverted u-shaped glass tube. which consists of salt mixture, like KCl and KNO_3 (NH_4Cl) and NH_4NO_3
- * The purpose of salt bridge is to connect anodic half cell and cathodic half cell reactions.
- * By using salt bridge . we prevent -the accumulation change of liquid junction potential.

Nernst Equation

In 1881, Walther Nernst proposed this equation to calculate the emf of the galvanic cells and oxidation potential, reduction potentials.

→ Let us consider a reversible reaction,



$$\text{Equilibrium constant } K = \frac{[\text{products}]}{[\text{reactants}]} = \frac{[M]}{[M^{+n}]} \quad (1)$$

According to statistical thermodynamics,

$$\text{Gibbs free energy } \Delta G = -nFE, \quad (1) \text{ definition}$$

$$\text{Standard Gibbs free energy } \Delta G^\circ = -nFE_0 \quad (2)$$

In eq(2), charge indicates the electrical workdone in reversible manner.

According to van't Hoff's equation

$$\Delta G = \Delta G^\circ + RT \ln K \quad (3)$$

Substitute (1) & (2) value in Eqn (3)

$$-nFE = -nFE_0 + RT \ln \frac{[M]}{[M^{+n}]} \quad (3)$$

Divide the above equation with "-nF" in both sides.

$$\frac{-nFE}{-nFE_0} = \frac{-nFE_0}{-nF} - \frac{RT}{-nF} \ln \frac{[M]}{[M^{+n}]}$$

$$E = E_0 - \frac{RT}{nF} \ln \frac{[M]}{[M^{+n}]}$$

$$E = E_0 + \frac{RT}{nF} \ln \frac{[M^{+n}]}{[M]}$$

The above equation is converted from natural logarithm into exponential logarithm. we get

$$E = E_0 + \frac{a \cdot 303RT}{nF} \log \frac{[M^{+n}]}{[M]}$$

$$R = \text{ideal gas constant} = 8.314 \text{ J/K}$$

$$T = \text{temperature} = 298 \text{ K}$$

$$F = \text{Faraday constant} = 96500 \text{ coulombs.}$$

Substitute R, T, F value we get,

$$E = E^{\circ} + \frac{0.0592}{n} \log \frac{[M^{+n}]}{[M]}$$

$$E = E^{\circ} + \frac{0.0592}{n} \log [M^{+n}] \quad [\because [M] = 1]$$

The Reduction potential formula,

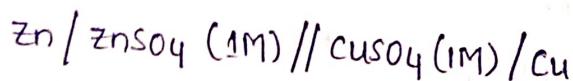
$$E = E_{\text{Red}}^{\circ} + \frac{0.0592}{n} \log [M^{+n}]$$

The oxidation potential formula,

$$E = E_{\text{Oxi}}^{\circ} - \frac{0.0592}{n} \log [M^{+n}]$$

Nernst Equation to the galvanic cell :-

Here cell notation (or) representation is



$$\text{EMF} = E_{\text{Cathode}} + E_{\text{Anode}} \quad [\because E_R - E_L]$$

$$E_{\text{cell}} = E_{\text{Cu}^{+2}} - E_{\text{Zn}^{+2}}$$

$$E = E^{\circ} + \frac{0.0592}{n} \log [M^{+n}]$$

$$E_{\text{Cu}^{+2}} = E_{\text{Cu}^{+2}}^{\circ} + \frac{0.0592}{2} \log [\text{Cu}^{+2}]$$

$$E_{\text{Zn}^{+2}} = E_{\text{Zn}^{+2}}^{\circ} + \frac{0.0592}{2} \log [\text{Zn}^{+2}]$$

$$E_{\text{cell}} = E_{\text{Cu}^{+2}}^{\circ} + \frac{0.0592}{2} \log [\text{Cu}^{+2}] - E_{\text{Zn}^{+2}}^{\circ} + \frac{0.0592}{2} \log [\text{Zn}^{+2}]$$

$$= E_{\text{Cu}^{+2}}^{\circ} - E_{\text{Zn}^{+2}}^{\circ} + \frac{0.0592}{2} \log \frac{[\text{Cu}^{+2}]}{[\text{Zn}^{+2}]}$$

$$\text{EMF} = E_R - E_L$$

$$= E_{\text{Cathode}} - E_{\text{Anode}}$$

Batteries :-

A device which converts chemical energy into electrical energy by spontaneous redox reactions. In reversible process and irreversible process is known as batteries.
⇒ on based upon construction of batteries these are divided into three types.

1. primary battery
2. secondary batteries
3. fuel batteries

1. primary batteries:-

⇒ A device which converts chemical energy into electrical energy by spontaneous redox reactions in irreversible process is known as primary batteries.
⇒ The another name of the primary batteries are non-rechargeable batteries.

Example:- Zn-air battery, Zn carbon cells.

2. Secondary batteries:

⇒ A device which converts chemical energy into electrical energy by spontaneous redox reaction in reversible process is known as secondary batteries.
⇒ The another name of secondary batteries are re-chargeable batteries.

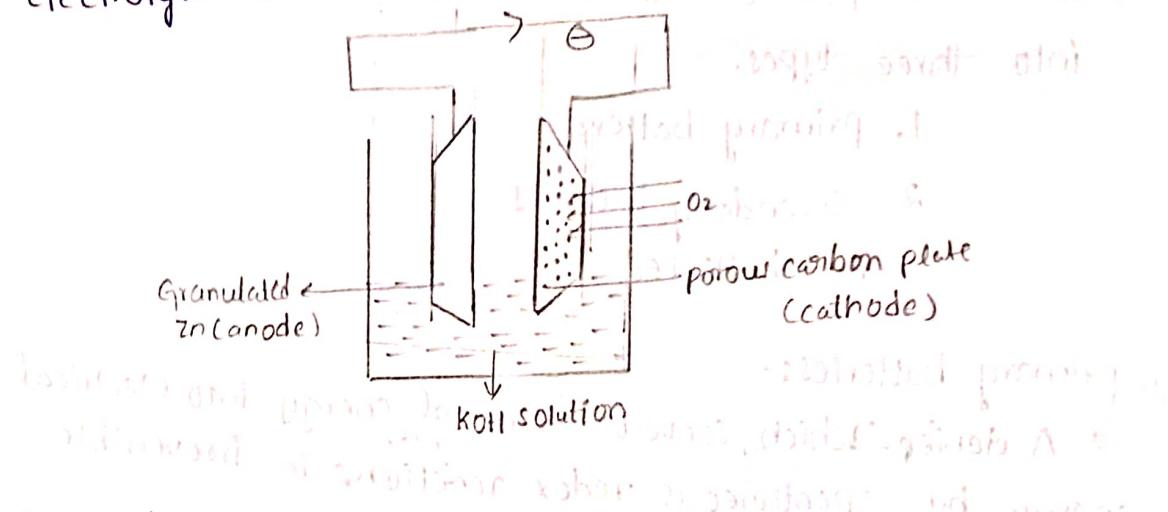
Example:- pb-acid battery & Li-ion battery

3. fuel batteries:-

⇒ A device which converts chemical energy into electrical energy into electrical energy by spontaneous redox reaction in the presence of fuel without combustion is known as "fuel batteries".

Example:- H₂-O₂ fuel cell, CH₃OH-O₂ fuel cell.

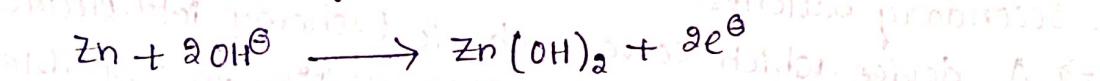
Zinc-air battery (Zn-air battery):
 The conversion of chemical energy into electrical energy in the presence of granulated zinc (anode) porous carbon plate (cathode) and KOH solution as an electrolyte is known as Zn-air battery.



At anode:-

The granulated zinc metal attracts OH^- ions from the electrolyte solution which forms zinc hydroxide and losing of electrons.

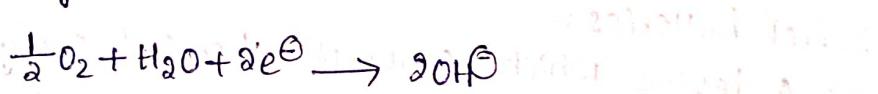
These lossing of electrons indicates oxidation reaction



At Cathode:-

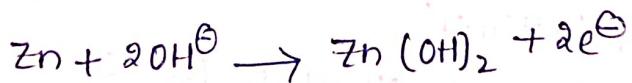
The porous carbon plate absorbs oxygen gas and water vapour from the atmosphere and gaining of electrons to form hydroxide ions.

Here, gaining of electrons indicates reduction reaction

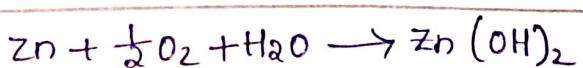
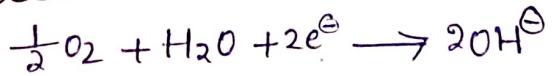


Cell reaction in Zn-air battery :-

At anode :-



At cathode :-

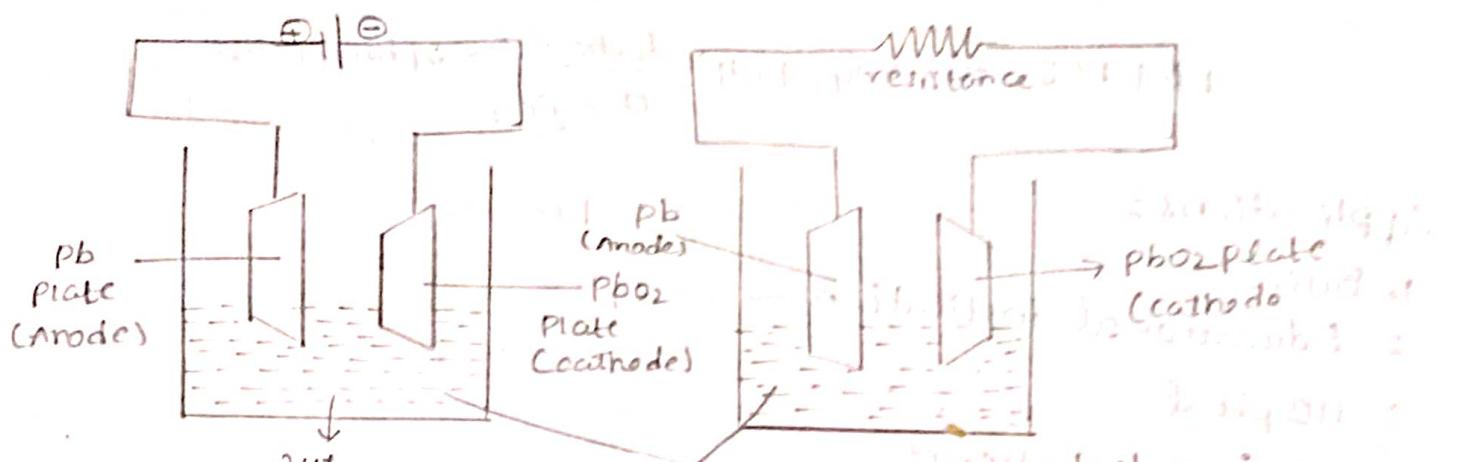


Applications :-

1. Wall clocks
2. Remotes
3. Torch lights
4. Radios.

Pb-acid battery:-

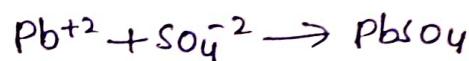
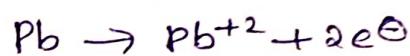
The conversion of chemical energy into electrical energy into electrical energy in the presence of sponge lead (anode), PbO_2 as cathode and 38% of H_2SO_4 solution as an electrolyte is known as Pb-acid battery.



At Anode:

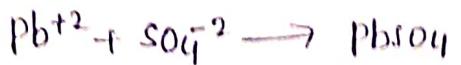
The sponge lead metal attracts SO_4^{2-} ions from the electrolyte solution to form lead ions (Pb^{+2}) and losing of electrons.

These losing of electrons indicates oxidation



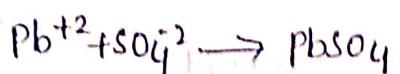
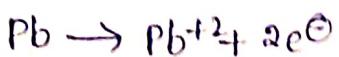
At Cathode:

The lead oxide plate attracts hydrogen ions from electrolyte and gaining of electrons to form lead ions and water. Again this lead ion attracts SO_4^{2-} ions from electrolyte which forms lead sulphate.

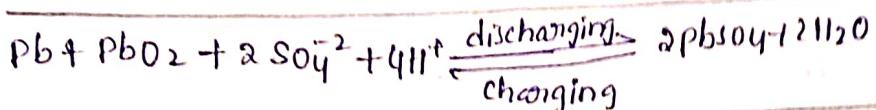
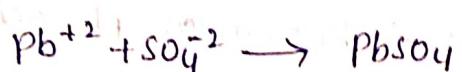


Cell reactions in pb-acid battery:

At Anode:



At cathode:

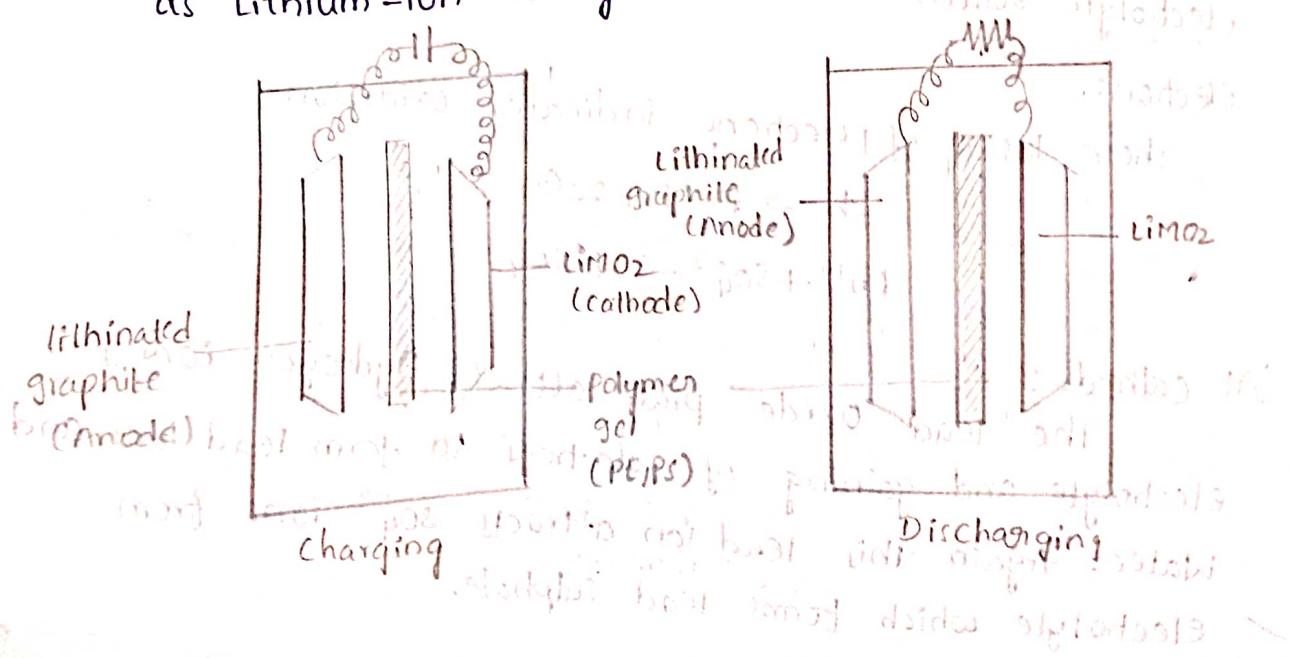


Applications:

1. Buses
2. Educational institutions
3. Hospital

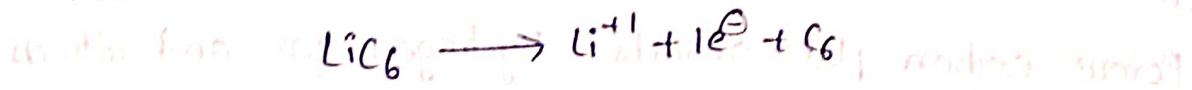
3. Lithium-ion battery:

The conversion of chemical energy into electrical energy in reversible process in the presence of lithiated graphite (anode), lithium metal oxide (LiMO_2) (cathode) and polymer gel acts as electrolyte is known as Lithium-ion battery.



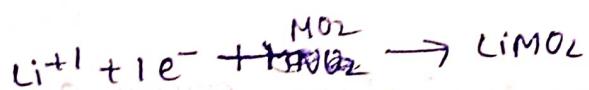
At Anode :-

Lithiated graphite electrode loses electron to form lithium ion and graphite. Here lossing of electrons indicates oxidation reaction.

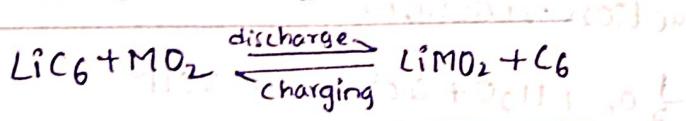
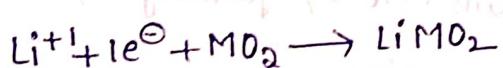
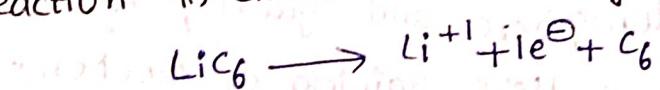


At cathode :-

At cathode Lithium ion gains one electron and also one Metal oxide to form lithium oxide. Here, Gaining of electrons indicates Reduction reaction.



Cell reaction in Lithium-ion battery:-

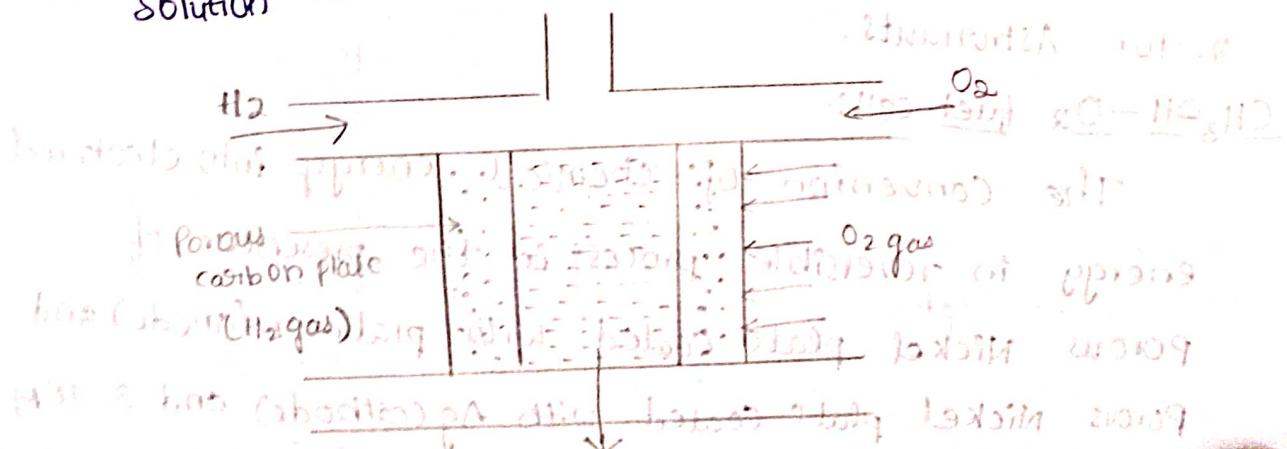


Applications :-

* Mobile battery.

~~H₂-O₂~~ fuel battery :-

The conversion of chemical energy into electrical energy in reversible process in the presence of porous carbon plates as both cathode and anode. and KOH solution as electrolyte is known as H₂-O₂ fuel battery.



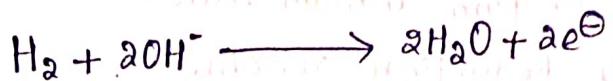
H₂ + O₂ → H₂O

Principle:- working of H₂-O₂ fuel cell.



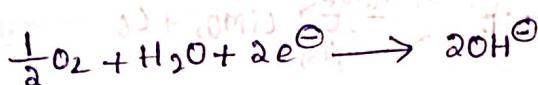
At anode:-

Porous carbon plate absorbs Hydrogen gas and attracts hydroxide ions and lossing of electrons. This lossing of electrons indicates oxidation reaction.

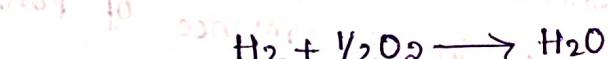
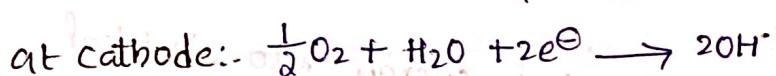
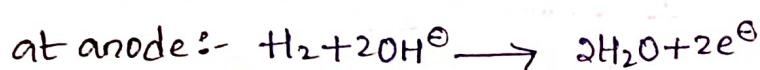


At cathode:-

Porous carbon plate absorbs oxygen gas and water vapour from the atmosphere and gaining of electrons to form hydroxide ions. Here, gaining of electrons indicates reduction reaction.



Cell reactions in H₂-O₂ fuel battery:-



Applications:-

1. used in space centers.

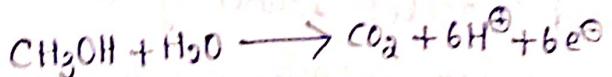
2. for Astronauts.

CH₃OH-O₂ fuel cell:-

The conversion of chemical energy into electrical energy in reversible process in the presence of porous Nickel plate coated with platinum (anode) and porous Nickel plate coated with Ag (cathode) and 3.75M of H₂SO₄ as an electrolyte is known as CH₃OH-O₂ fuel cell.

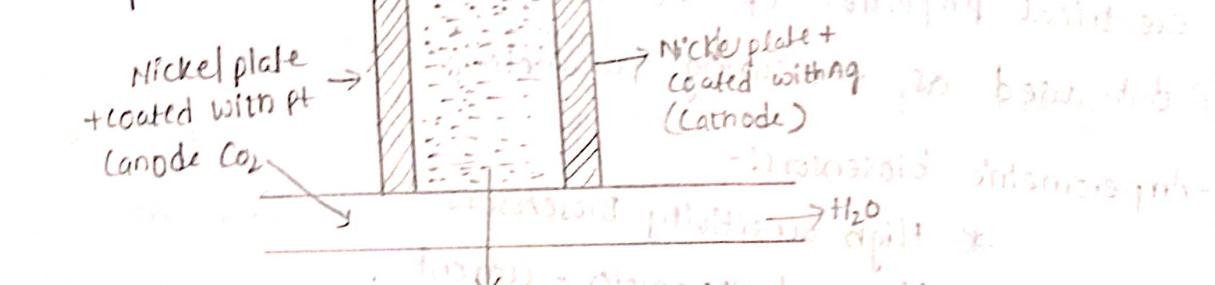
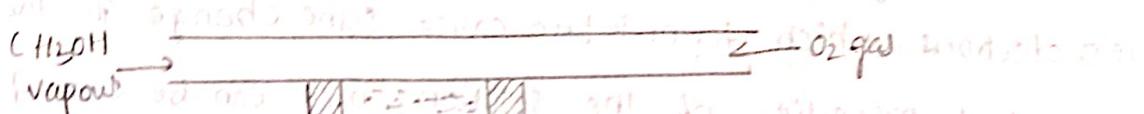
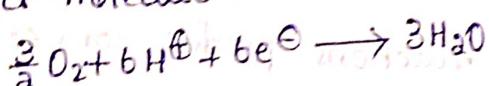
At Anode:-

Methyl alcohol interacts with water molecule from the electrolyte solution to produce carbon dioxide, Hydrogen ions and losing of $6e^-$. Here, lossing of electrons indicates oxidation reactions.

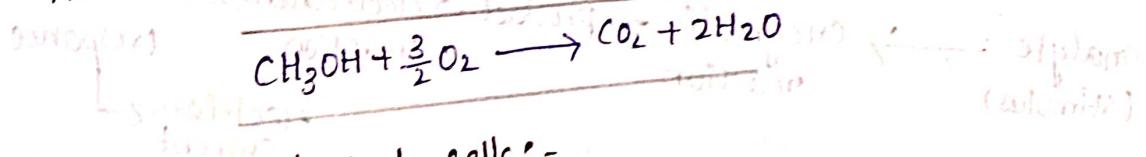
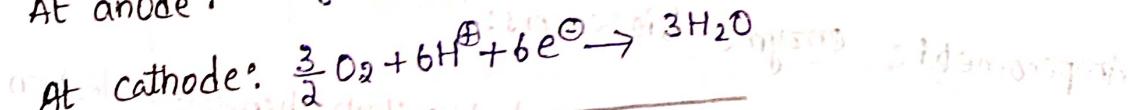
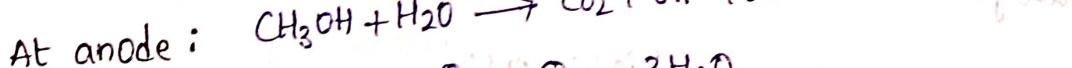


At cathode:-

The oxygen gas interacts with hydrogen ions from the electrolyte solution because of H_2SO_4 is strong acid which produce more no. of H^+ ions and gaining of electrons to form water molecule



Cell reactions in $CH_3OH - O_2$ fuel cell.



Advantages of fuel cells:-

1. No noise pollution

2. No thermal pollution

3. Easily transportation.

- 4) Easily maintainence
- 5) The bi-products also useful for astronauts i.e water etc.
- 6) Eco-friendly Energy source

Applications :-

* These are mainly used in rocket launching stations like Space satellite and research purpose

Electrochemical biosensor :-

Other names immuno-sensors, chemical canaries, biochips, bio-computers.

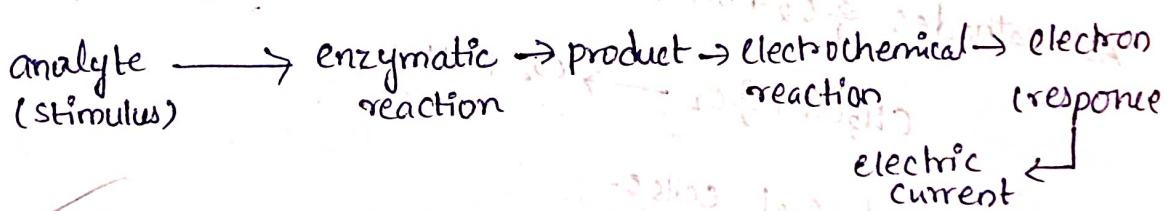
Principle :-

Many chemical reactions produce (or) consume ions (or) electrons which inturn cause some change in the electrical properties of the solution which can be sensed out used as measuring parameters.

Amperometric biosensors :-

- * High sensitivity Biosensors
- * Measured parameter - current
- * The simplest amperometric biosensors in common usage involve the clark oxygen electrode.

Amperometric enzyme biosensors :-



Amperometric Glucose Biosensor :-

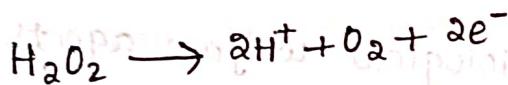
- Developed by updike and flicks.
- Enzyme glucose oxidise, catalyze the oxidation of Glucose by molecular oxygen producing gluco lactone and hydrogen peroxide

→ GOx requires a redox co-factor - Flavin Adenine dinucleotide (FAD)



Hydrogen peroxide is oxidized at a platinum electrode.

The no. of electron transferred at electrode surface is directly proportional to the number of glucose molecules present in the blood.

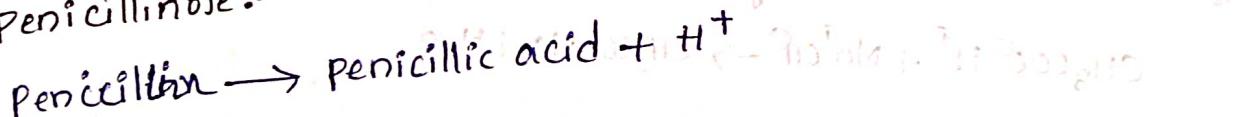


Potentiometric sensors:-

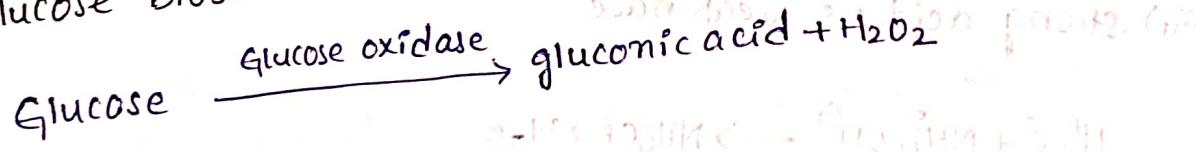
- * Working principle - When ramp voltage is applied to an electrode in a solution, a current flow occurs because of electrochemical reactions.
- * Measured parameters - oxidation/reduction potential of an electrochemical reaction.

Example:-

1. Penicilliose:-



2. Glucose biosensors:-



3. Urea biosensors:-

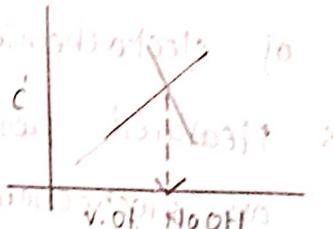
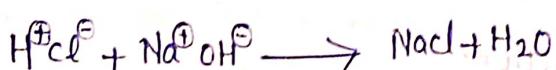


Applications:-

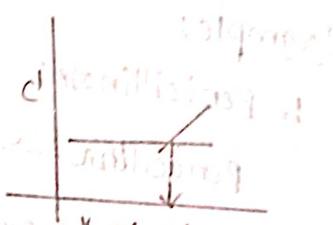
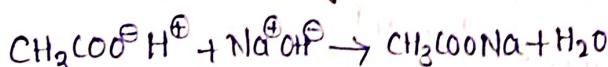
1. Food Analysis
2. Study of Biomolecules & their Interaction
3. Drug development
4. Crime Detection
5. Medical Diagnosis
6. Environmental field monitoring
7. Quality control
8. Industrial process control
9. Detection system for Biological warfare agents.
10. Manufacturing of pharmaceuticals & Replacement organs.
11. Detection of viral, fungal, Bacterial disease of plants.

Conductometric titration of Acid vs Base:-

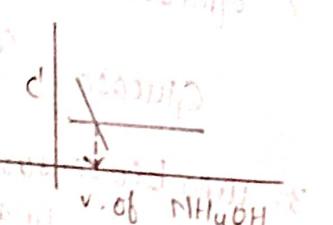
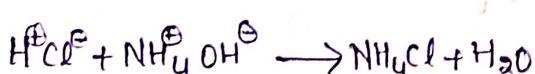
(i) strong acid vs strong base.



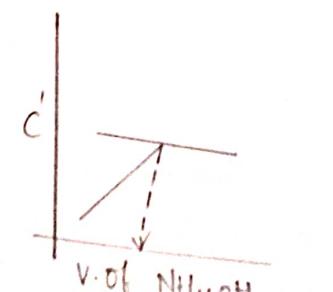
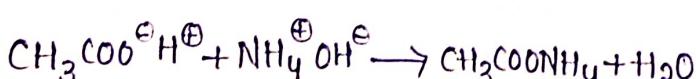
(ii) weak acid vs strong base



(iii) strong acid vs weak base



(iv) weak acid vs weak base



concept of conductivity:-

Ohm's law (or) Resistance Specific resistance

Resistance is directly proportional to the length between the electrodes and inversely proportional to cross sectional area of the electrodes

$$R \propto l/a$$

$$\text{units cm} \quad R = \rho \frac{l}{a}$$

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

where R = Resistance ; units are ohm

ρ = Specific resistance

units are $\Omega = \text{ohm} \frac{\text{cm}^2}{\text{cm}}$

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

conductance :- The reciprocal of electrical resistance is known as

conductance.

$$C = 1/R$$

units of $C = 1/\text{ohm}$

The conductance or mho = ohm^{-1} (or) Mho (or) Siemens.

specific conductance :-

The reciprocal of specific resistance is known as

specific conductance.

\Rightarrow The specific conductance is designated as Kappa (κ)

$$\kappa = 1/\rho$$

$$\text{units} \Rightarrow \rho = \frac{1}{\text{ohm.cm}}$$

$$\kappa = \text{ohm}^{-1}\text{cm}^{-1}$$

Molar conductance:
The strength of ions producing from one mole of electrolyte solution is known as Molar conductance.

$$\Lambda_{\text{molar}} = \frac{1000 k_0}{c}$$

units = $\Lambda_{\text{molar}} = \frac{\text{cm}^3 \cdot \text{ohm}^{-1} \cdot \text{cm}^{-1}}{\text{moles}}$

$$= \text{mole}^{-1} \cdot \text{cm}^2 \cdot \text{ohm}^{-1}$$

Equivalent conductance:

The strength of ions producing from one equivalent weight of the electrolyte solution is known as equivalent conductance.

⇒ The Equivalent conductance is designated as $\Lambda_{\text{equivalent}}$

$$\Lambda_{\text{equivalent}} = \frac{1000 k_0}{c}$$

$$\text{units} = \Lambda_{\text{equivalent}} = \frac{\text{cm}^3 \cdot \text{ohm}^{-1} \cdot \text{cm}^{-1}}{\text{Equivalence}}$$

$$= \text{Equivalence}^{-1} \cdot \text{cm}^2 \cdot \text{ohm}^{-1}$$

Cell constant:

The Ratio of length between the electrodes and cross section area of the electrode is known as cell constant.

⇒ Cell constant is designated as '(x)'

$$x = l/a$$

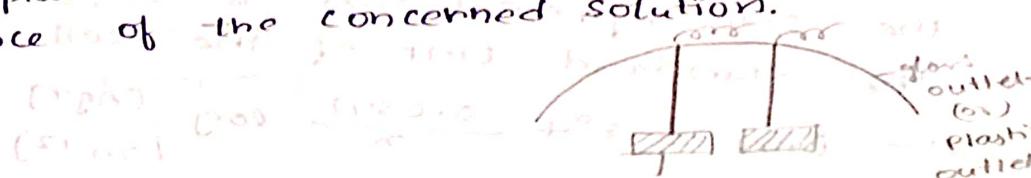
$$\text{unit} = x = \text{cm} / (\text{cm}^2)$$

$$= \frac{\text{cm}}{\text{cm}^2} = \text{cm}^{-1}$$

conductivity cell

The two platinum foil plates which are coated with platinum black powder covered with glass outlet with plastic is known as conductivity cell.

(or) plastic But left is connected to the power supply
 \Rightarrow In this platinum plates connected to the concerned solution.
 to give conductance of the concerned solution.



problem based upon EMF & Nernst equation.

$$EMF = E_{\text{cathode}} - E_{\text{anode}}$$

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

Nernst equation

$$E = E^{\circ} + \frac{0.0591}{n} \log \frac{[\text{Red}]}{[\text{Oxid}]}$$

1. calculate the emf of the following reaction at 25°C where the cell representation is $\text{Cu}^{+2} + \text{Zn} \rightleftharpoons \text{Zn}^{+2} + \text{Cu}$ where

$$E_{\text{Zn}^{+2}} = -0.763 \text{ V} \quad \text{and} \quad E_{\text{Cu}^{+2}} = +0.337 \text{ V}$$

$$\begin{aligned} \text{Sol:-- } EMF &= E_{\text{cathode}} - E_{\text{anode}} \\ &= E_{\text{Cu}^{+2}} - E_{\text{Zn}^{+2}} \\ &= 0.337 - (-0.763) \end{aligned}$$

$$EMF = 1.1 \text{ V}$$

2. what is the single electrode potential of half cell for zinc electrode dipping in a 0.01 M zinc solution

$$E^{\circ} = 0.763 \text{ V}$$

$$\begin{aligned} \text{Sol:-- Given} \\ \text{Molarity (M+n)} &= 0.01 \text{ M} \\ EMF &= E^{\circ} - \frac{0.0592}{n} \log [\text{ions}]_{\text{initial}}$$

$E^\circ = 0.763 - 0.0591 \log \frac{P_{O_2}}{P_{O_2}^0}$

3. What is the EMF of the following cell at 25°C
- The cell representation is $Zn/Zn^{+2}(0.2M) \parallel Ag(Ag^{+})/Ag$
 - The standard EMF of the cell is $E^\circ = 1.54V$

$$\text{Sol: } E^\circ = E^\circ + \frac{0.0591}{2} \log \frac{[Ag^+]}{[Zn^{+2}]}$$

$$= 1.54 + \frac{0.0591}{2} \log \frac{[0.002]}{[0.2]}$$

$$= 1.54 - 0.139 = 1.401V$$

Smp Potentiometric Sensors

Smp Conductometry (Lab)

4. calculate the EMF of cell representation of electrochemical cell
 $Pb^{+2} + Cd \rightarrow Pb + Cd^{+2}$ given that $E^\circ_{cell} = 0.277V$ $[Cd^{+2}] = 0.02M$ $[Pb^{+2}] = 0.2M$.

$$EMF = E^\circ_{cell} + \frac{0.0591}{2} \log \frac{\text{Oxidation}}{\text{Reduction}}$$

$$= 0.277 - \frac{0.0591}{2} \log \frac{[Cd^{+2}]}{[Pb^{+2}]}$$

$$= 0.277 - \frac{0.0591}{2} \log \frac{[0.02]}{[0.2]}$$

Why Li-ion has future :-

low cost

more power

longer range

faster charging time

greater flexibility and

Improved safety over their inhouse contemporaries.

Limitations of conductometric titrations.

1. Only a few specific redox titrations can be done with the help of this process.
2. The accuracy of conductometric titration is low when the concentrations of electrolyte are high, making the titration process unsatisfactory.