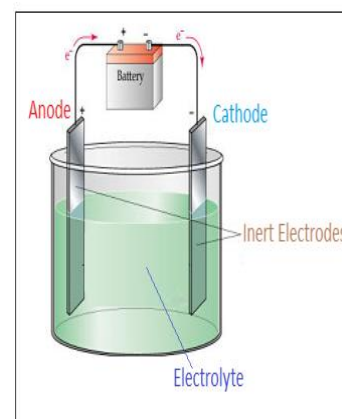


Unit-I Electrochemistry and Battery Chemistry

Electrochemical cells:

Electrolytic cell:

The cell which converts electrical energy into chemical energy is called as electrolytic cell. An electrolytic cell consists of two metallic or graphite rods dipping in solution of electrolyte and connected to a battery. The rod connected to a positive terminal of the battery acts as anode while that connected to negative terminal acts as cathode. It is to be noted that in an electrolytic cell anode is of positive sign (+), but the cathode is of negative sign (-).

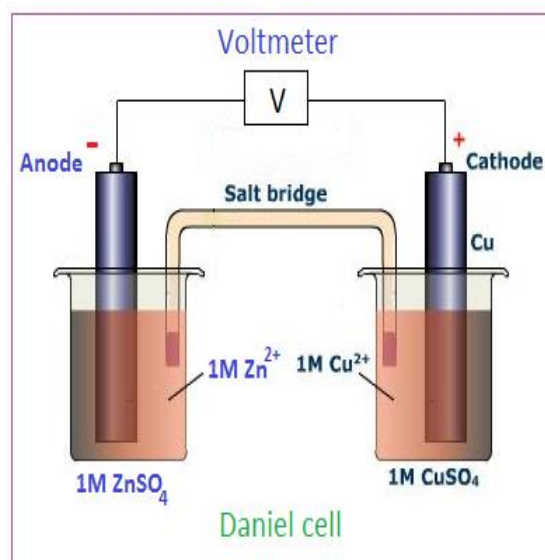


Galvanic Cell:

Galvanic cell is a device which converts chemical energy into electrical energy by means of a spontaneous redox reaction. It consists of two electrodes immersed in the solution of their respective electrolytes. The two electrodes called, half cells, are connected by a salt bridge. The electrodes are connected to potentiometer or voltmeter by an electrical wire. The electrode at which oxidation occurs is called as anode (Negative (-) Electrode). The electrode at which reduction occurs is

called as Cathode (Positive (+) Electrode).

Eg: Daniel cell



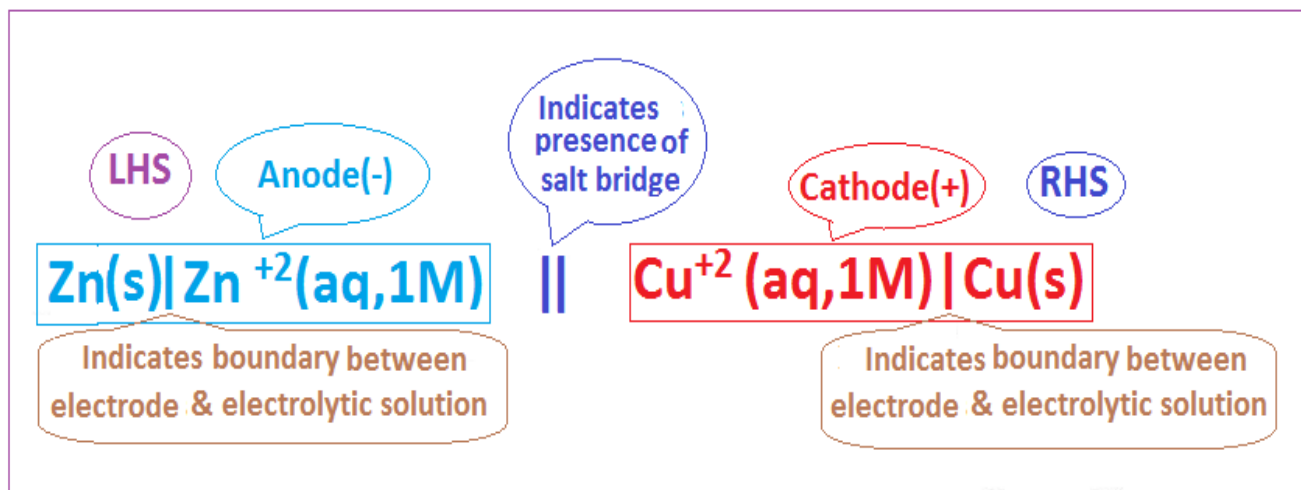
Salt bridge & Its Role:

Salt bridge is an inverted U-tube consists of a concentrated solution of electrolytes like KCl, KNO₃, and NH₄Cl. These electrolytes are preferred because the mobilities of cation and anion are almost equal.

[Type here]

1. It connects the solutions of two half-cells and completes the cell circuit.
2. It prevents diffusion of the solutions from one half-cell to the other
3. It keeps the solutions in two half-cells electrically neutral.
4. It prevents liquid-liquid junction-potential.

Representation of cell Or Cell Notation:



Difference between Electrolytic cell and Galvanic Cell

Electrolytic Cell	Galvanic cell
Electrical energy (from an external source) is converted into chemical energy.	Chemical energy (of a redox reaction) is converted into electrical energy.
The electrodes may be made of same or different metals.	The electrodes are made of dissimilar metals.
At the electrodes, non-spontaneous reaction take place only when the electrical energy is supplied .	At the electrodes , spontaneous redox reaction take place.
Anode – Positive ; Cathode - Negative	Anode – Negative ; Cathode – Positive
Ions are discharged on both the electrodes	Ions are discharged only on the cathode.
Both electrodes are inserted in the same electrolytic solution.	Each electrode is dipped in electrolytic solution of its own ions.

[Type here]

EMF of a cell:

When two electrodes are combined current flows from higher potential to lower potential, the potential difference which is responsible for the flow of current is called EMF of the cell.

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

Where, E_{Cathode} and E_{Anode} are reduction potentials of cathode and anode respectively.

Electrode potential:

When a metal is placed in its electrolytic solution, the metal acquires either positive or negative charge with respect to electrolytic solution which results a potential difference is developed at the interface between electrode and electrolytic solution. This potential difference is called electrode potential. Or

Tendency of an electrode to lose or gain electrons when it is in contact with its own ionic solution is called as electrode potential.

The tendency of an electrode to lose electrons i.e. to get oxidised, is called Oxidation potential, whereas, the tendency of an electrode to gain electrons i.e. to get reduced, is called Reduction potential.

Standard electrode potential:

The potential of an electrode, at a given temperature, depends upon the concentration of solution of its ions.

"The tendency of a pure metal or pure gas at 1 atm pressure to undergo oxidation or reduction when it is in contact with 1M solution of its ions at 25°C, is called as standard electrode potential".

SRP: *It is the tendency of a pure metal or pure gas at 1 atm pressure to undergo reduction when it is in contact with 1M solution of its ions at 25°C.*

SOP: *It is the tendency of a pure metal or pure gas at 1 atm pressure to undergo oxidation when it is in contact with 1M solution of its ions at 25°C.*

Electrochemical series:

Electrochemical series is the arrangement of electrodes from top to bottom in the increasing order of their standard reduction potentials, which are determined with reference to standard Hydrogen electrode. It is also called as electromotive or activity series.

[Type here]

B.E I Sem (Autonomous)
Chemistry notes (Group -B: IT, EEE, Mechanical, Automobile) (2021-22)

Electrode	Electrode Reaction	Standard reduction potential, E^0 volts
Li^+/Li	$\text{Li}^+ + \text{e}^- \longrightarrow \text{Li}$	-3.04
K^+/K	$\text{K}^+ + \text{e}^- \longrightarrow \text{K}$	-2.93
Na^+/Na	$\text{Na}^+ + \text{e}^- \longrightarrow \text{Na}$	-2.71
Mg^{+2}/Mg	$\text{Mg}^{+2} + 2\text{e}^- \longrightarrow \text{Mg}$	-2.37
Al^{+3}/Al	$\text{Al}^{+3} + 3\text{e}^- \longrightarrow \text{Al}$	-1.66
Zn^{+2}/Zn	$\text{Zn}^{+2} + 2\text{e}^- \longrightarrow \text{Zn}$	-0.76
Fe^{+2}/Fe	$\text{Fe}^{+2} + 2\text{e}^- \longrightarrow \text{Fe}$	-0.44
Ni^{+2}/Ni	$\text{Ni}^{+2} + 2\text{e}^- \longrightarrow \text{Ni}$	-0.23
Sn^{+2}/Sn	$\text{Sn}^{+2} + 2\text{e}^- \longrightarrow \text{Sn}$	-0.14
Pb^{+2}/Pb	$\text{Pb}^{+2} + 2\text{e}^- \longrightarrow \text{Pb}$	-0.13
$\text{H}^+, \text{H}_2/\text{Pt}$	$2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2$	0
$\text{Sn}^{+4}/\text{Sn}^{+2}$	$\text{Sn}^{+4} + 2\text{e}^- \longrightarrow \text{Sn}^{+2}$	0.14
$\text{AgCl}, \text{KCl}/\text{Ag}$	$\text{AgCl} + \text{e}^- \longrightarrow \text{Ag} + \text{Cl}^-$	0.222
$\text{Hg}_2\text{Cl}_2, \text{KCl}, \text{Hg}/\text{Pt}$	$\text{Hg}_2\text{Cl}_2 + 2\text{e}^- \longrightarrow 2\text{Hg} + 2\text{Cl}^-$	0.28
Cu^{+2}/Cu	$\text{Cu}^{+2} + 2\text{e}^- \longrightarrow \text{Cu}$	0.34
$\text{OH}^-, \text{O}_2, \text{H}_2\text{O}/\text{Pt}$	$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \longrightarrow 4\text{OH}^-$	0.401
$\text{I}^-, \text{I}_2/\text{Pt}$	$\text{I}_2 + 2\text{e}^- \longrightarrow 2\text{I}^-$	0.535
$\text{Q}, \text{H}^+, \text{QH}_2/\text{Pt}$	$\text{Q} + 2\text{H}^+ + 2\text{e}^- \longrightarrow \text{QH}_2$	0.6992
$\text{Fe}^{+3}, \text{Fe}^{+2}/\text{Pt}$	$\text{Fe}^{+3} + \text{e}^- \longrightarrow \text{Fe}^{+2}$	0.77
Ag^+/Ag	$\text{Ag}^+ + \text{e}^- \longrightarrow \text{Ag}$	0.80
$\text{O}_2, \text{H}^+, \text{H}_2\text{O}/\text{Pt}$	$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \longrightarrow 2\text{H}_2\text{O}$	1.23
$\text{Cr}_2\text{O}_7^{2-}, \text{H}^+, \text{Cr}^{+3}/\text{Pt}$	$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \longrightarrow 2\text{Cr}^{+3}$	1.33
$\text{Cl}^-, \text{Cl}_2/\text{Pt}$	$\text{Cl}_2 + 2\text{e}^- \longrightarrow 2\text{Cl}^-$	1.36
$\text{MnO}_4^-, \text{H}^+, \text{Mn}^{+2}/\text{Pt}$	$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \longrightarrow \text{Mn}^{+2}$	1.51
$\text{F}^-, \text{F}_2/\text{Pt}/$	$\text{F}_2 + 2\text{e}^- \longrightarrow 2\text{F}^-$	2.87

[Type here]

Applications of electrochemical Series:

1. Relative strengths of Oxidising and reducing agents

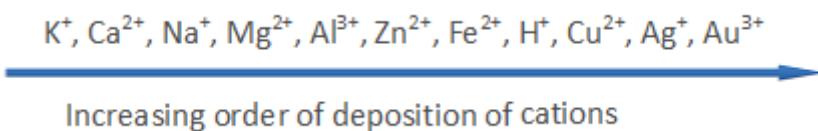
The electrode with low SRP acts as strong reducing agent while the electrode with high SRP acts as strong oxidizing agent.

we can observe, from the electrochemical series, that lithium acts as the strongest reducing agent, Fluorine acts as the strongest oxidising agent. When two electrodes are combined, the electrode with **higher potential undergoes reduction** i.e., **acts as cathode**.

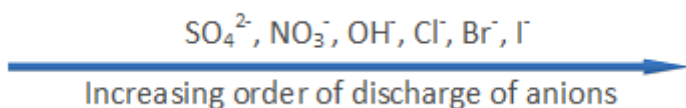
2. Predicting the product of electrolysis:

In case two or more types of positive and negative ions are present in solution, during electrolysis certain ions are deposited or liberated at the electrodes in preference to others.

In general, if there is competition between two or more than two cations, the cation with high value standard reduction potential (Stronger Oxidising Agent) is discharged first at the cathode.



Similarly, if there is competition between two or more than two anions the anion with low value of standard reduction potential (Stronger Reducing Agent) is liberated first at the anode.



3. Tendency of metals to displace hydrogen from dilute acids

The tendency of metals to displace hydrogen from dilute acids can be predicted i.e., the reactivity of metals with dilute acids to liberate hydrogen can be predicted. The

metals with negative potential can displace hydrogen from dilute acids.

4. Feasibility of Redox reaction

A redox reaction to be spontaneous E_{cell} should be +ve and ΔG must be -ve. It can be predicted by calculating E_{cell} from electrochemical series.

5. Calculation of equilibrium constant of a Redox reaction

Equilibrium constant can be calculated if E_{cell}^0 is known (E_{cell}^0 Can be calculated from electrochemical series).

We have,

$$\Delta G^0 = -nFE_{\text{cell}}^0 \quad \& \quad \Delta G^0 = -RT \ln K$$

$$RT \ln K = nFE_{\text{cell}}^0$$

$$\ln K = \frac{nF}{RT} E_{\text{cell}}^0 \qquad \text{Log } K = \frac{2.303nF}{RT} E_{\text{cell}}^0$$

Nernst Equation & Its derivation:

The Nernst Equation helps us to determine cell potential or electrode potential under non-standard and non-equilibrium conditions. It relates the measured cell potential/electrode potential to the reaction quotient.

Consider a redox reaction, $aA + bB \xrightleftharpoons{Q} cC + dD$

$$Q = \frac{[C]^d [D]^d}{[A]^a [B]^b}$$

The decrease in the free energy of such equilibrium is given as, $-\Delta G = -\Delta G^0 - 2.030RT \log Q$

When a system is operating reversibly the work done by the system is equal to the decrease in the free energy of the system.

If the charge associated with a redox process is “ nF coulombs” and potential generated is E Volts, the net electrical work done by the reversible cell is “ nFE Joules”

$$-\Delta G^0 = nFE_{\text{cell}}^0 \quad \& \quad -\Delta G = nFE_{\text{cell}}$$

$$nFE_{\text{cell}} = nFE_{\text{cell}}^0 - RT \ln Q$$

$$E_{\text{cell}} = E_{\text{cell}}^0 - \frac{RT}{nF} \ln Q$$

[Type here]

$$E_{cell} = E_{cell}^0 - \frac{2.303RT}{nF} \log Q$$

$$E_{cell} = E_{cell}^0 - \frac{2.303RT}{nF} \log \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

At $T = 25^\circ\text{C} = 298\text{ K}$, R - Gas Constant = 8.314 J/mole , F – Faraday = 96500 Coulombs

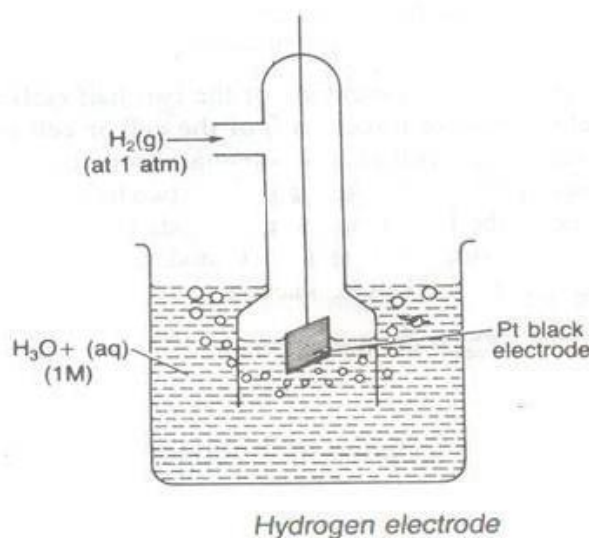
$$\text{Therefore, } E_{cell} = E_{cell}^0 - 0.0591 \log \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Applications of Nernst Equation:

1. Nernst equation is used to study the effect of electrolyte concentration on electrode potential.
2. Nernst equation can be used to calculate cell potentials / Electrode potentials under non-standard and non-equilibrium condition.
3. Determination of unknown concentration of one of the ionic species involved in a redox reaction is possible with help of Nernst equation provided E_{cell}^0 and concentration of other ionic species is known.
4. The pH of a solution can be calculated from the measurement of EMF and Nernst equation.
5. It can also be used to find the valency (n) of reactive species.
6. It can be used to calculate equilibrium constant of redox reaction.

Types of Electrodes:

1. Gas Electrode (Standard Hydrogen Electrode):



[Type here]

Hydrogen electrode is the primary standard electrode. It consists of a small Platinum strip coated with Platinum black as to adsorb H₂ gas. Platinum is welded to the Pt strip and sealed in a glass tube as to make contact with the outer circuit. The Platinum strip and glass tube is surrounded by an outer glass tube which has an inlet for H₂ gas at top and a no. of holes at the base for the escape of excess of H₂ gas. The Platinum strip placed in an acid solution which has H⁺ ions of 1M concentration. Pure hydrogen gas passed through the inlet at 1 atmospheric pressure. A part of gas is adsorbed and the rest escapes through holes. This gives equilibrium between the adsorbed H₂ and H⁺ ions in the solution.

The temperature of the cell is maintained at 25 °C. By international agreement the H₂ electrode is assigned a potential of 0.0 volts.

The electrode represented as Pt, H₂ (g) / H⁺ (aq). Hydrogen electrode can acts as anode or cathode w.r.to the other electrode to which it is connected.

Anode: Oxidation: H₂ (g) → 2H⁺ + 2e⁻
(E_{ele}=0.0 V)

Cathode: Reduction: 2H⁺ + 2e⁻ → H₂ (g)

(E_{ele} = 0.0 V)

$$E_{H^+/H_2} = E^0_{H^+/H_2} - \frac{0.059}{2} \log \frac{P_{H_2}}{(H^+)^2}$$

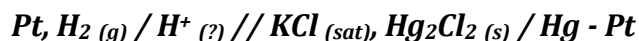
$$E_{H^+/H_2} = 0.0591 \log H^+$$

$$E_{H^+/H_2} = -0.0591 \text{pH}$$

Thus, potential of hydrogen electrode depends on pH of the solution.

It is used to measure pH of solutions.

For example, if the H₂ electrode is coupled with Saturated Calomel Electrode (SCE),



$$E_{\text{cell}} = E_{\text{SCE}} - E_{\text{H}_2}$$

$$= 0.242 - (-0.0591 \text{pH})$$

$$E_{\text{cell}} = 0.242 + 0.0591 \text{pH}$$

$$\text{pH} = \frac{E_{\text{cell}} - 0.242}{0.0591}$$

Type equation here.

Drawbacks:

1. Its construction is difficult i.e. it is difficult to maintain the concentration of [Type here]

H⁺ ions 1M and pressure 1 atm.

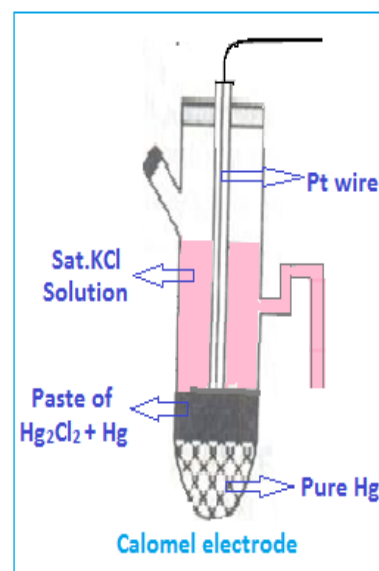
2. Cannot be used in the presence of ions of many metals.
3. It cannot be used in solutions containing Redox systems.

2. Metal – insoluble salt electrode (Calomel Electrode):

It consists of a pure metal in contact with one of its sparingly soluble salts and a solution of a soluble salt having a common anion with the sparingly soluble salt.

Construction and working of Calomel electrode:

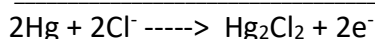
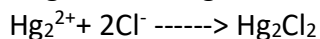
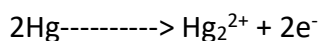
Calomel electrode is an example of metal-metal insoluble salt electrode, which consists of a glass tube containing two side arms. One of the arms is to connect to salt bridge and the other one is to introduce the components of electrode. Pure Mercury is placed at the bottom of the glass tube. Mercury is covered with a paste of Mercury and Mercurous chloride, over which a solution of potassium chloride is placed. A platinum wire sealed in a thin glass tube, dipping into the mercury layer, is used for making electrical contact with the external circuit.



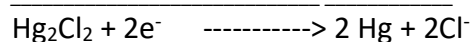
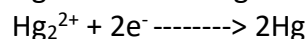
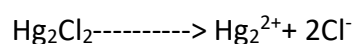
Calomel electrode is represented as **Pt-Hg/ Hg₂Cl₂/KCl (aq).**

Calomel electrode acts as either anode or cathode w.r.to the other electrode to which it is connected.

If it acts as anode, it involves oxidation.



If it acts as cathode it involves reduction.



When it acts as anode, mercury is converted to Mercurous ions. Mercurous ions would combine with Cl⁻ ions furnished by KCl to

[Type here]

form sparingly soluble Hg_2Cl_2 which results in the fall in the concentration of Cl^- ions. When it acts as cathode, Hg_2^{2+} ions, furnished by the sparingly soluble Hg_2Cl_2 , would be discharged at the electrode and there is an increase in the concentration of Cl^- ions. Thus, Calomel electrode is said to be reversible to Cl^- ions.

The reduction potential of calomel electrode is given by,

$$E_{CE} = E_{CE}^0 - \frac{0.0591}{2} \log \frac{[\text{Hg}]^2 [\text{Cl}^-]^2}{[\text{Hg}_2\text{Cl}_2]}$$

$$E_{CE} = 0.280\text{V} - 0.0591 \log [\text{Cl}^-]$$

Since $[\text{Hg}] = [\text{Hg}_2\text{Cl}_2] = 1$ and $E_{CE}^0 = 0.280\text{ V}$

The potential of the calomel electrode depends upon the concentration of the chloride ion. If the concentration of KCl is 0.1 N, 1.0 N and saturated, calomel electrode is called as Decinormal, Normal and Saturated calomel electrode and their electrode potentials at 25°C are 0.333V, 0.28 V and 0.242V respectively.

Saturated KCl	0.2415 V
1.0N KCl	0.2800 V
0.1N KCl	0.3338 V

Advantages: Its construction is very easy, potential measurements are reproducible.

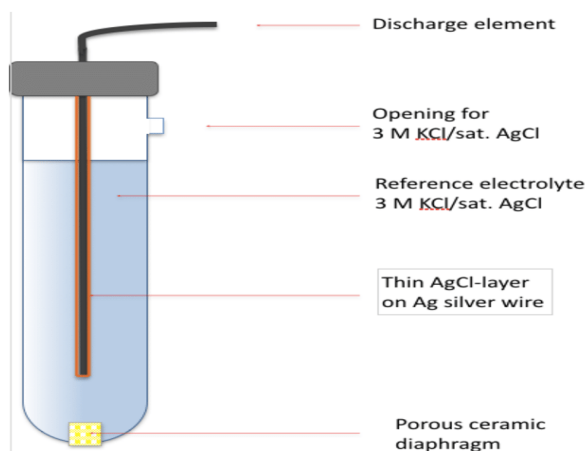
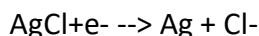
Disadvantages: Since Hg_2Cl_2 breaks at 50°C , it can't be used above this temperature.

Silver-silver chloride electrode: This is another widely used reference electrode. It is reversible and stable and can be combined with cells containing chlorides without inserting liquid junctions.

Silver chloride is deposited electrolytically on a silver or platinum wire and it is then immersed in a solution containing chloride ions. Its standard electrode potential with respect to the standard hydrogen electrode is 0.2224 V at 298 K. the electrode is represented as:



And at the electrode the reaction is

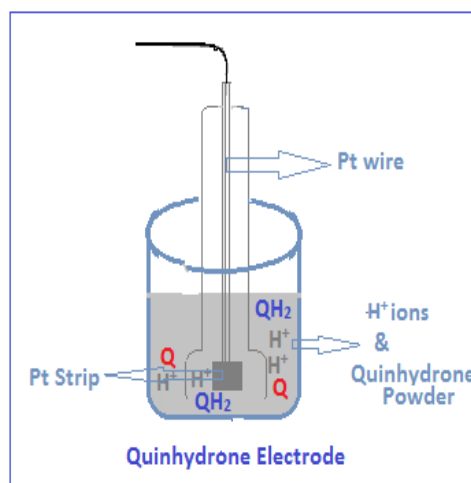


3.Redox Electrodes(Quinhydrone Electrode):

A Redox electrode consists of an inert material like Pt or Au dipped into a solution containing solution of a chemical species in two different oxidation states.

Construction and working of Quinhydrone electrode:

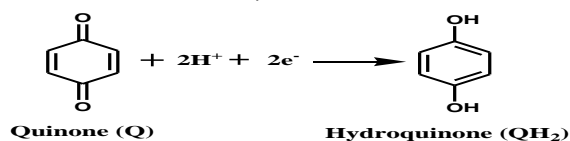
The Quinhydrone electrode is an example of redox electrode, which is constructed by adding a pinch of Quinhydrone powder, which is an equimolar mixture of quinone and hydroquinone to an acid solution, under test, taken in a beaker. An inert Pt electrode is dipped into the solution for making electrical contact.



The electrode is represented as $\text{H}^+(\text{aq}); \text{Q}; \text{QH}_2/\text{Pt}$.

It will either act as anode or cathode depending on the other electrode to which it is connected.

If it acts as cathode, then the electrode reaction is,



The electrode potential is calculated by using Nernst equation,

[Type here]

$$E_{QHE} = E_{QHE}^0 - \frac{0.0591}{2} \log \frac{[QH_2]}{[Q][H^+]^2}$$

$$E_{QHE} = E_{QHE}^0 + \frac{0.0591}{2} \log [H^+]^2 \quad (\text{Since } [QH_2] = [Q] = 1)$$

$$E_{QHE} = E_{QHE}^0 + 0.0591 \log [H^+]$$

$$E_{QHE} = E_{QHE}^0 - 0.0591 \text{pH}$$

$$E_{QHE} = 0.699V - 0.0591 \text{pH} \quad (\text{On Hydrogen Scale } E_{QHE}^0 = 0.699V)$$

Potential of Quinhydrone electrode depends on pH of the solution with which it is in contact. Therefore, Quinhydrone electrode can also be used to measure the pH of solution.

Advantages:

The electrode setup is very easy, pH values obtained are accurate, small quantities of solution are sufficient for the measurement.

Disadvantages:

It cannot be used in the solutions that would react with quinone or hydroquinone.

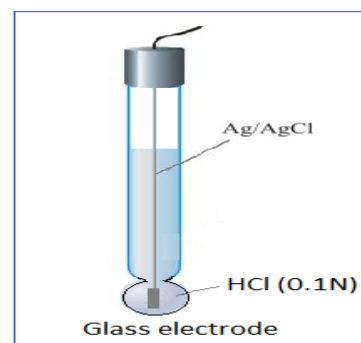
Since it is a weak acid, cannot be used above pH 8.5 at which the dissociation of Quinhydrone becomes appreciable.

It is oxidized by air in strongly alkaline medium.

(5) Ion selective electrode(Glass electrode):

Ion-selective electrode possesses the ability to respond only to certain specific ions, thereby developing a potential with respect to that specific ions in a solution and ignoring the other ions totally. In other words, the potential developed by an ion – selective electrode depends only on the concentration of ions of interest.

Glass electrode is made up of a special type of glass having low melting point and high electrical conductance. The bottom of the glass electrode is blown in the form a bulb, which is filled with a 0.1 N HCl solution to provide constant hydrogen ion concentration. A silver wire coated with AgCl is inserted into the bulb for making electrical contact.



The representation of glass electrode is, $\text{Ag/AgCl/0.1N HCl/Glass membrane}$. The potential of the glass electrode is given as, $E_G = E_G^0 - 0.0591 \text{pH}$

Determination of pH of a solution using quinhydrone electrode coupled with saturated calomel electrode :

The electrode potential of Quinhydrone electrode, which is reversible to H^+ ions, depends on pH of a solution; therefore it is used to determine pH of a solution. To determine pH of a solution, quinhydrone electrode is combined with a reference electrode like calomel electrode. The cell consists of Quinhydrone and Calomel electrode is represented as,

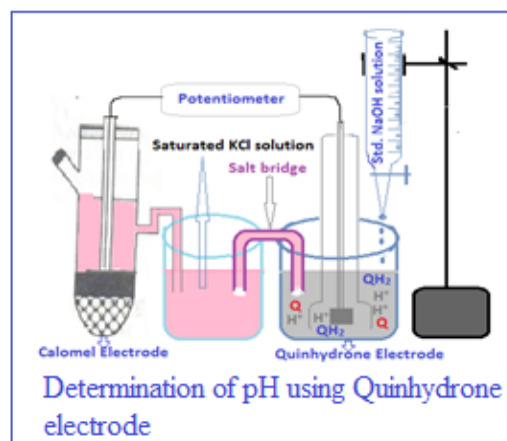
$\text{Pt, Hg/ Hg}_2\text{Cl}_2, \text{Cl}^-(\text{Sat, aq}) // \text{H}^+(\text{aq}), \text{Q, QH}_2/\text{Pt}$.

$$E_{\text{cell}} = E_{\text{QHE}} - E_{\text{SCE}}$$

$$E_{\text{cell}} = 0.699\text{V} - 0.0591 \text{pH} - .242\text{V}$$

$$E_{\text{cell}} = 0.457 - 0.0591 \text{pH}$$

$$\text{pH} = \frac{0.457 - E_{\text{cell}}}{0.0591}$$



Problems Based on Applications of Nernst Equations:

- Consider the cell, $\text{Zn/Zn}^{2+}(\text{aq}) // \text{Cu}^{2+}(\text{aq})/\text{Cu}$. The standard electrode potentials are;



- Write down the cell reaction
- Calculate the EMF of the cell.

Solution: (i) Reduction potential of Zn is less than Cu, hence Zn acts anode and Cu acts as cathode.



$$(ii) E_{\text{Cell}}^0 = E_{\text{Cu}^{2+}/\text{Cu}}^0 - E_{\text{Zn}^{2+}/\text{Zn}}^0$$

$$= 0.35\text{V} - (-0.763\text{V})$$

[Type here]

$$= 1.113 \text{ V}$$

2. Calculate the electrode potential at a copper electrode dipped in a 0.1M solution of copper sulphate at 25°C. The standard electrode potential of Cu^{+2}/Cu system is 0.34V at 298K.

Ans: 0.31045 Volts

3. E_{cell}^0 of a given cell, $\text{Zn(s)} / \text{Zn}^{+2} \text{ 1M(aq)} // \text{Cu}^{+2} \text{ 1M (aq)} / \text{Cu(s)}$, is 1.10 V at 25°C. Calculate equilibrium constant for the cell reaction, $\text{Zn(s)} + \text{Cu}^{+2}(\text{aq}) \rightleftharpoons \text{Cu(s)} + \text{Zn}^{+2}(\text{aq})$.

Ans : 1.67×10^{37}

4. A zinc rod is placed in a 0.1M solution of ZnSO_4 at 25°C. Calculate the potential of the electrode at this temperature, assuming 96% dissociation of ZnSO_4 and $E_{\text{Zn}^{+2}/\text{Zn}}^0 = -0.76 \text{ V}$.

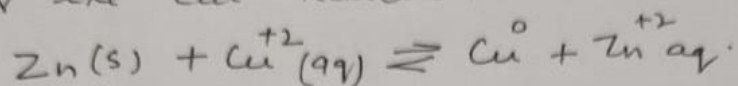
Ans: $E_{\text{Zn}^{+2}/\text{Zn}} = -0.79 \text{ V}$.

5. The emf of a cell corresponding to the reaction, $\text{Zn} + 2\text{H}^+ \rightarrow \text{Zn}^{+2} + \text{H}_2(\text{g})$ 1 atm is 0.28V at 25°C. Write the half cell reactions and calculate the pH of the solution at hydrogen electrode. ($E_{\text{Zn}^{+2}/\text{Zn}}^0 = -0.76 \text{ V}$ $E_{\text{H}^+/\text{H}_2}^0 = 0\text{V}$)

Ans : pH= 8.6218

Problem -1

E°_{cell} of a given cell, $\text{Zn(s)}|\text{Zn}^{+2}(\text{1M})||\text{Cu}^{+2}(\text{1M})|\text{Cu}$ is 1.10V at 25°C . calculate equilibrium constant for the cell reaction



Solution : given $E^\circ = 1.10\text{V}$, $n=2$, $F=96500$
 $T=298$, $R=8.314\text{ J/K}$

$$-\Delta G^\circ = RT \ln K$$

and also

$$-\Delta G^\circ = nFE^\circ$$

or

$$RT \ln K = nFE^\circ$$

$$\ln K = \frac{nFE^\circ}{RT}$$

$$\ln K = \frac{2 \times 96500 \times 1.10}{8.314 \times 298}$$

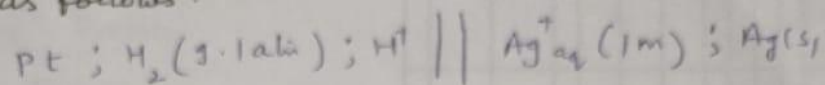
$$\log K = \frac{2 \times 96500 \times 1.1}{2.303 \times 8.314 \times 298}$$

$$\log K = 37.207$$

$$K = \text{antilog } 37.207$$

$$K = 1.61 \times 10^{37}$$

Soln Problem: An electrochemical cell is represented as follows:

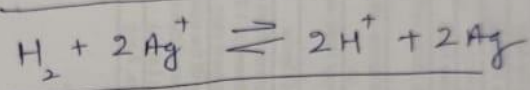
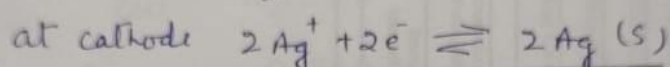
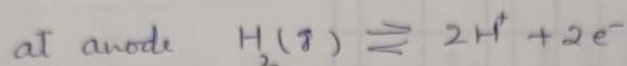


given that

given $E_{\text{cell}} = 0.900 \text{ V}$ at 25°C ; $E_{\text{Ag}^+/\text{Ag}}^\circ = 0.80 \text{ V}$

Find the pH of the solution

Sol: Reaction involved



According to Nernst equation

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.059}{n} \log \frac{[\text{H}^+]^2}{[\text{Ag}^+]^2}$$

$$(E_{\text{cell}}^\circ = 0.80 - 0.0)$$

substituting the values

$$0.90 = 0.80 - \frac{0.059}{2} \log \frac{[\text{H}^+]^2}{(1)^2}$$

$$0.90 - 0.80 = -0.0295 \log [\text{H}^+]^2$$

$$\frac{0.1}{0.0295} = -\log [\text{H}^+]^2$$

$$3.3898 = -2 \log [\text{H}^+]$$

$$\boxed{1.6949} = -\log [\text{H}^+] \approx \text{pH}$$

