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**Subject: Chemistry**

**Class: Semester-2**

**Paper: C3T: Inorganic Chemistry**

**Topic: Redox and Precipitation Reaction**

**PART 3**

**Comments:** Go through the whole lesson as much as possible for conception. But some topics such as “**standard electrode potential, Nerst equation and calculation of EMF of some cells, calculation of free energy, equilibrium constant and application of standard redox potential values**” are very important for preparation of examination. Also complete the assignment given below.

**[N.B.- Acknowledgement of indebtedness to Mr. Sibshankar Das, my respected Teacher regarding collection of study materials in Inorganic Chemistry]**

## • Galvanic cell

The driving force of a rxn can be measured in terms of EMF of a Galvanic cell. Galvanic cell is a device in which oxidation and reduction rxn, takes place in two separate containers and electrical energy is produced during this rxn. It is an arrangement of two electrodes and electrolyte in two separate containers. This chemical energy is converted into electrical energy in a Galvanic cell.

When such a cell is connected in series with an ammeter and the circuit is closed, there occurs a deflection in the ammeter indicating a flow of current through the circuit due to existence of a pot.

potential difference between the two electrodes. This potential difference which is responsible for the <sup>flow</sup> of current from the electrode of higher potential to the one of the lower potential, is known as the electromotive force or EMF of the cell and is expressed in volts.

### • Cell Rxn and EMF

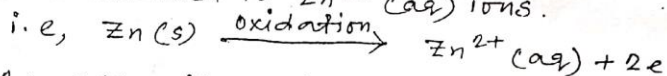
Let us explain the working of a Galvanic cell by taking the example of Daniell cell. This cell consists of two beakers, one of which contains 1(M) solution of  $ZnSO_4$  and the other beaker contains 1(M) solutions of  $CuSO_4$ . A zinc rod is dipped into  $ZnSO_4$  sol<sup>n</sup> while a Cu rod is dipped into  $CuSO_4$  solution. These metallic rods are called electrodes.

The electrodes are connected with an ammeter. A key is also inserted in the circuit. The solutions of the two beakers are called electrolyte solutions. These solutions are connected together by an inverted U tube containing saturated solution of an inert electrolyte like  $KCl$ ,  $KNO_3$  etc. The U tube is called salt bridge.

When the circuit is completed by inserting the key in the circuit, the needle of the ammeter shows a deflection. The deflection of the needle indicates that electric current is flowing in the external circuit.

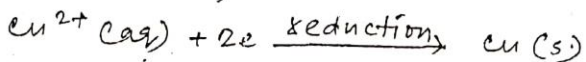
Following observations are made,

i)  $Zn(s)$  is oxidised to  $Zn^{2+}(aq)$  ions.



As a result,  $Zn^{2+}$  ion obtained by the oxidation of Zn, go into the solution of  $ZnSO_4$  and hence the concentration of  $Zn^{2+}$  ions in  $ZnSO_4$  sol<sup>n</sup> increases. Due to the conversion of Zn into  $Zn^{2+}$  ions, Zn rod loses its weight.

ii) Electrons released at Zn electrodes towards Cu electrodes through the outer circuit. These electrons are accepted by  $Cu^{2+}$  ions of  $CuSO_4$  sol<sup>n</sup>. Which are reduced to Cu. i.e;



As a result, since  $Cu^{2+}$  ions of  $CuSO_4$  sol<sup>n</sup> are reduced to Cu, the concentration of  $Cu^{2+}$  ions in  $CuSO_4$  sol<sup>n</sup> decreases. Cu obtained by reduction of  $Cu^{2+}$  ions gets deposited on Cu rod and hence Cu rod gains weight.



### • Single electrode Potential

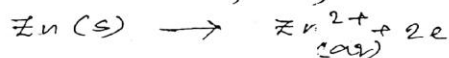
Let us consider, a metal: metal ion electrode which is obtained by immersing a metal rod in its salt solution i.e; sol<sup>n</sup> containing metal cations. In this electrode there exist a separation of charges between the metal and its salt solution. This separation of charges results that a net negative or positive charge is developed on the metal with respect to its solution. The negative or positive charge developed on the metal produces electrical potential difference between the metal and its solution. This electrical potential difference is called single electrode potential. Thus single electrode potential is a measure of the tendency of an electrode in a half cell to gain or loss electrons.

### • Origin of electrode Potential

How the electrode potential is generated, has been explained by Nernst. By this theory which is called Nernst Theory, Let us consider the following two examples.

#### a) Zn metal plate placed in a sol<sup>n</sup> containing Zn<sup>2+</sup> ions.

When a Zn metal is placed in a sol<sup>n</sup> containing Zn<sup>2+</sup> ions, Zn atoms loss electrons and are oxidised to Zn<sup>2+</sup> ions, i.e;



Electrons thus generated get accumulated on the Zn metal and hence make the Zn plate negatively charged. In order to neutralise the negative charge created on the surface of the metal plate, a layer of Zn<sup>2+</sup> ions is formed round the surface of Zn plate. In this way an electrical layer is generated between the metal plate and the solution. This electrical layer developed a definite potential difference between the metal and metal ions of the sol<sup>n</sup>. This potential difference is called electrode potential of Zn(s) / Zn<sup>2+</sup>(aq) electrode.

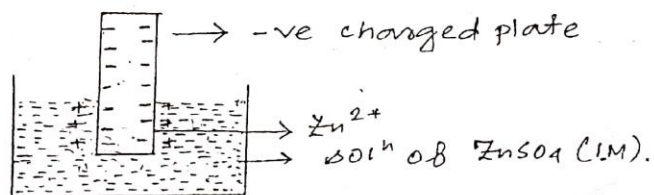


Fig: origin of single electrode potential of Zn(s) / Zn<sup>2+</sup>(aq) etc.

b) Cu metal plate placed in a sol<sup>n</sup> containing Cu<sup>2+</sup>

When a Cu metal plate is placed in a solution containing Cu<sup>2+</sup> ions, Cu plate becomes positively charged and in order to neutralise the positive charge on the plate, a layer of sulphate ions is formed round the surface of the positively charged plate. In this case an electrical layer is generated between the metal plate and the solution, containing Cu<sup>2+</sup> ions. Thus electrical layer gives rise to a potential difference between the metal and metal ions in solution. This potential difference is called electrode potential of Cu(s) / Cu<sup>2+</sup>(aq) electrode.

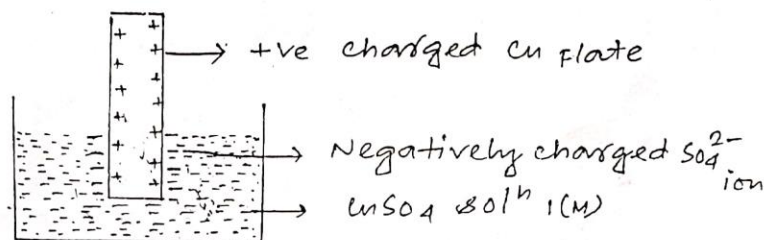
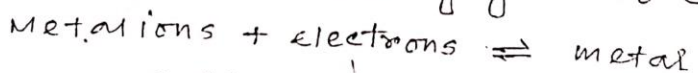


Fig: Origin of single electrode potential of Cu<sup>2+</sup>(aq) / Cu(s) electrode.

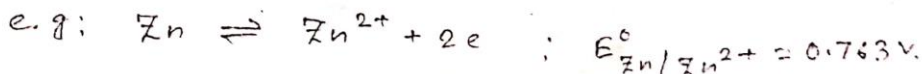
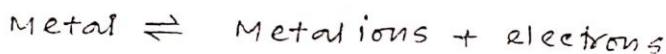
\* standard electrode potential

The standard electrode potential of a metal-metal ion electrode is the EMF of a cell made by coupling a normal hydrogen electrode with the metal-metal ion electrode obtained by immersing the metal in its salt solutions of 1 (M) concentration at 25°C. The potential of normal H electrode is taken as zero.

The standard reduction potential gives a measure of the tendency of metal ions (at unit concentration) to get reduced by gain of electrons.



The standard oxidation potential measured the capacity of a metal to lose electrons in a sol<sup>n</sup> of its ions at unit concentration.



Reduction potential = - Oxidation potential

### • Sign Convention.

According to the present convention, the value of electrode potential is expressed as reduction electrode potential. Thus if we say that the value of standard electrode potential of  $Zn(s) / Zn^{2+}(aq)$  electrode is  $-0.763$  volt, this means that the given value is the value of standard reduction potential of  $Zn$  electrode.

$$\text{i.e.; } E_{Zn^{2+}/Zn}^{\circ} = -0.763 \text{ volt}$$

### • Negative Reduction potential and positive Reduction Potential.

According to the present convention the value of the electrode potential is expressed as reduction electrode potential. This reduction potential may have a negative value or positive value; depending on whether the given electrode acts as an anode or as a cathode of the galvanic cell obtained by coupling the given electrode with a normal hydrogen electrode (NHE or SHE). When the given electrode acts as an anode, the electrode has a negative value for its standard reduction potential. On the other hand when the given electrode acts as a cathode, the electrode has a positive value for its standard reduction potential.

If two electrodes are coupled together to get a galvanic cell, then the electrode having less value of its standard reduction potential will constitute the anode of the resulted cell and the electrode having more value of its reduction potential will constitute the cathode of the cell.

### • Standard EMF of a Galvanic cell.

When all reactants and the products involved in the cell reaction of a galvanic cell are at unit concentration (1M), the EMF of the cell is called standard EMF of the cell which is represented as  $E_{cell}^{\circ}$ . It is also called standard cell potential.

### • Rel<sup>n</sup> between free energy and cell EMF

We know that 1 equivalent of any substance corresponds to 1 Faraday or F. Coulombs of electricity. Suppose  $n$  equivalents of the reactants in a chemical cell are present per mole. Then for

each mole of the reactant converted to products,  $nF$  Coulombs of electric charge pass through the cell. If  $E$  is the EMF developed by the cell rxn, the net electrical work available would be  $nFE$  Joule. When the cell operates, this work is equal to the decrease in free energy per mole of the reactant. Denoting this by  $- \Delta G$ , we obtain the relation,

$$- \Delta G = nFE \quad \text{or} \quad \Delta G = - nFE$$

For thermodynamic principles, we know that at constant temp<sup>r</sup> and pressure  $\Delta G$  is -

- i) Negative for any spontaneous process.
- ii) Positive for any non-spontaneous process.
- iii) Equal to zero for any system at equilibrium.

Since,  $\Delta G = - nFE$ , we may readily conclude that the EMF of cell will be,

- i) Positive when the cell rxn is spontaneous.
- ii) Negative when the cell rxn is non-spontaneous.
- iii) Equal to zero when the cell rxn reaches equ<sup>m</sup>.

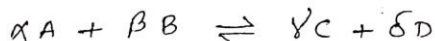
At standard state, i.e; when the activities of all species appearing in electrode rxn are unity.

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

Where,  $\Delta G^\circ$  = standard free energy change  
 $E^\circ$  = standard EMF of the cell.

- Derive Nernst equation for an electrochemical cell and establish the relation between standard free energy change and equilibrium constant for a given cell rxn.

Let us consider an electrochemical cell rxn



where the activities of all species undergo change. For species A free energy change for a mole is

$$\begin{aligned} \Delta G_A &= G_A + \alpha RT \ln a_A \\ &= G_A^\circ + RT \ln a_A^\alpha \end{aligned}$$

Similarly the expression for species B, C and D are

$$\begin{aligned} \Delta G_B &= G_B^\circ + RT \ln a_B^\beta \\ \Delta G_C &= G_C^\circ + RT \ln a_C^\gamma \\ \Delta G_D &= G_D^\circ + RT \ln a_D^\delta \end{aligned}$$



The change in free energy for the overall cell rxn is expressed as

$$\Delta G = \sum G_{\text{product}} - \sum G_{\text{reactant}}$$

$$= \Delta G^\circ + RT \ln \frac{a_C^\gamma \cdot a_D^\delta}{a_A^\alpha \cdot a_B^\beta} \quad \text{--- (i)}$$

Where  $\Delta G^\circ$  = standard free energy change.

We know  $\Delta G = -nFE$  and  $\Delta G^\circ = -nFE^\circ$ , where  $E^\circ$  is the standard EMF of the cell. Thus eqn (i) becomes

$$-nFE = -nFE^\circ + RT \ln \frac{a_C^\gamma \cdot a_D^\delta}{a_A^\alpha \cdot a_B^\beta}$$

$$\Rightarrow E = E^\circ - \frac{RT}{nF} \ln \frac{a_C^\gamma \cdot a_D^\delta}{a_A^\alpha \cdot a_B^\beta} \quad \text{--- (ii)}$$

This is known as Nernst equation.  
From equation (ii)

$$E = E^\circ - \frac{2.303 RT}{nF} \log \frac{a_C^\gamma \cdot a_D^\delta}{a_A^\alpha \cdot a_B^\beta}$$

$$= E^\circ - \frac{0.059}{n} \log \frac{a_C^\gamma \cdot a_D^\delta}{a_A^\alpha \cdot a_B^\beta} \quad \text{at } 25^\circ\text{C.}$$

This is also the another form of Nernst equation.

$$\text{Where } \frac{2.303 RT}{F} = 0.059 \quad \text{at } 25^\circ\text{C.}$$

For dilute sol<sup>n</sup> equation (ii) can be written as

$$E = E^\circ - \frac{RT}{nF} \ln \frac{c_C^\gamma \cdot c_D^\delta}{c_A^\alpha \cdot c_B^\beta} \quad \text{--- (iii)}$$

At equilibrium, since the cell could not perform any work as EMF will be zero, hence equation (iii) becomes

$$E^\circ = \frac{RT}{nF} \ln K_{eq} \quad [K_{eq} = \text{Equilibrium Constant}]$$

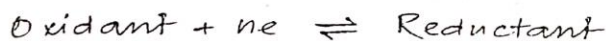
$$\Rightarrow nFE^\circ = RT \ln K_{eq}$$

$$\Rightarrow -\Delta G^\circ = RT \ln K_{eq}$$

$$\Rightarrow \boxed{\Delta G^\circ = -RT \ln K_{eq}} \quad \text{--- (iv)}$$

This is the relation between standard free energy change and equilibrium constant for a given cell rxn.

\* For a rxn

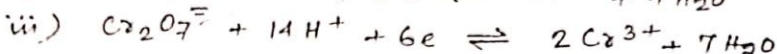
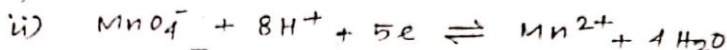
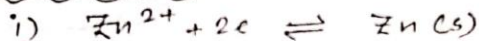


The Nernst equation can be written as

$$E = E^\circ - \frac{RT}{nF} \ln \frac{[\text{Reductant}]}{[\text{Oxidant}]}$$

$$= E^\circ + \frac{RT}{nF} \ln \frac{[\text{Oxidant}]}{[\text{Reductant}]}$$

• Write down the Nernst equations for the following equations.



i) For the rxn  $\text{Zn}^{2+} + 2e \rightleftharpoons \text{Zn}(s)$  The Nernst equation can be written as,

$$E_{\text{Zn}^{2+}/\text{Zn}(s)} = E^\circ_{\text{Zn}^{2+}/\text{Zn}(s)} + \frac{RT}{nF} \ln \frac{a_{\text{Zn}^{2+}}}{a_{\text{Zn}(s)}}$$

Since Zn is in the solid state, its activity is 1. Hence

$$E_{\text{Zn}^{2+}/\text{Zn}(s)} = E^\circ_{\text{Zn}^{2+}/\text{Zn}(s)} + 2.303 \frac{RT}{nF} \log a_{\text{Zn}^{2+}}$$

For dilute sol<sup>n</sup> and at 25°C. This equation can be written as,

$$E_{\text{Zn}^{2+}/\text{Zn}(s)} = E^\circ_{\text{Zn}^{2+}/\text{Zn}(s)} + \frac{0.059}{2} \log c_{\text{Zn}^{2+}}$$

Here  $n=2$  and  $\frac{2.303RT}{F} = 0.059$  at 25°C.

$E^\circ_{\text{Zn}^{2+}/\text{Zn}(s)}$  = Standard reduction potential of Zn/Zn<sup>2+</sup> electrode.

ii) The given equation is  $\text{MnO}_4^- + 8\text{H}^+ + 5e \rightleftharpoons \text{Mn}^{2+} + 4\text{H}_2\text{O}$   
The Nernst equation can be written as

$$E_{\text{MnO}_4^-/\text{Mn}^{2+}} = E^\circ_{\text{MnO}_4^-/\text{Mn}^{2+}} + \frac{RT}{nF} \ln \frac{a_{\text{MnO}_4^-} \cdot a_{\text{H}^+}^8}{a_{\text{Mn}^{2+}} \cdot a_{\text{H}_2\text{O}}^4}$$

For dilute sol<sup>n</sup> and at 25°C, This equation can be written as

$$E_{\text{MnO}_4^-/\text{Mn}^{2+}} = E^\circ_{\text{MnO}_4^-/\text{Mn}^{2+}} + \frac{RT}{5F} \ln \frac{[\text{MnO}_4^-] [\text{H}^+]^8}{[\text{Mn}^{2+}]}$$

Since  $a_{\text{H}_2\text{O}} = 1$ .

For one molar H<sup>+</sup> ion sol<sup>n</sup>, This equation can be written as

$$E_{\text{MnO}_4^-/\text{Mn}^{2+}} = E^\circ_{\text{MnO}_4^-/\text{Mn}^{2+}} + \frac{0.059}{5} \log \frac{[\text{MnO}_4^-]}{[\text{Mn}^{2+}]}$$

Here  $\frac{2.303RT}{F} = 0.059$  at 25°C and  $n=5$ .

iii) For the rxn  $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightleftharpoons 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$ , the Nernst equation can be written as;

$$E_{\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}} = E^\circ_{\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}} + \frac{RT}{nF} \ln \frac{a_{\text{Cr}_2\text{O}_7^{2-}}}{a_{\text{Cr}^{3+}}^2} \quad \left[ \text{Since } a_{\text{H}_2\text{O}} = 1 \right]$$

$$\therefore E_{\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}} = E^\circ_{\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}} + \frac{0.059}{6} \log \frac{a_{\text{Cr}_2\text{O}_7^{2-}}}{a_{\text{Cr}^{3+}}^2}$$

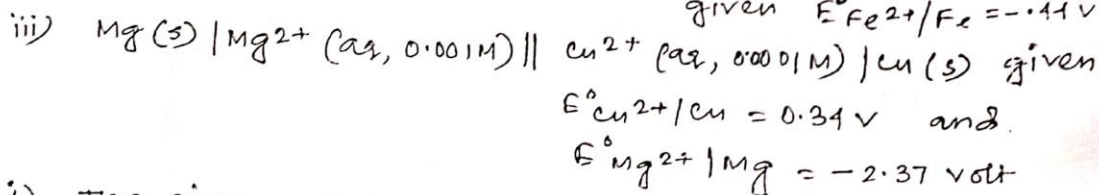
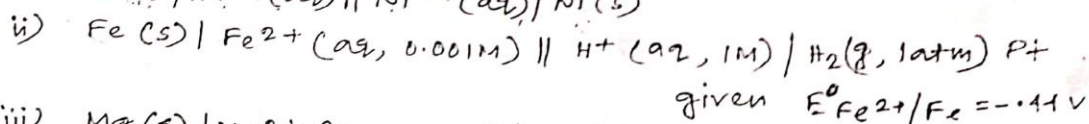
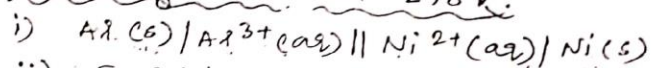
For dilute solution and at  $25^\circ\text{C}$ , this equation can be written as -

$$E_{\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}} = E^\circ_{\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}} + \frac{0.059}{6} \log \frac{C_{\text{Cr}_2\text{O}_7^{2-}}}{C_{\text{Cr}^{3+}}^2}$$

Here  $n=6$  and  $\frac{2.303RT}{F} = 0.059$  at  $25^\circ\text{C}$ .

$E^\circ_{\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}}$  = standard reduction potential of  $\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}$  electrode.

• Write Nernst equation and calculate EMF of the following cells at 298 K.



i) The given cell composed of two electrodes which are  $\text{Al}(s) | \text{Al}^{3+}(aq)$  [anode] and  $\text{Ni}^{2+}(aq) | \text{Ni}(s)$  [cathode] rxns, taking place at two electrodes are

Oxidation rxn at anode;  $[\text{Al}(s) \rightarrow \text{Al}^{3+}(aq) + 3\text{e}^-] \times 2$

Reduction rxn at cathode  $[\text{Ni}^{2+}(aq) + 2\text{e}^- \rightarrow \text{Ni}(s)] \times 3$

Total cell rxn;  $2\text{Al}(s) + 3\text{Ni}^{2+}(aq) \rightleftharpoons 2\text{Al}^{3+}(aq) + 3\text{Ni}(s)$

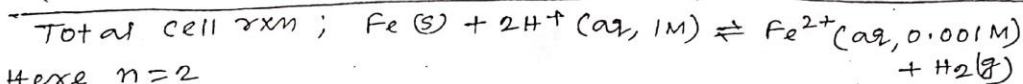
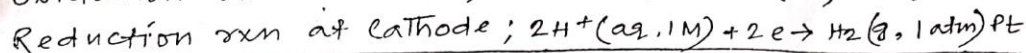
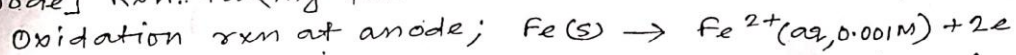
Nernst equation for this rxn at 298K is written as

$$E_{\text{cell}} = E^\circ_{\text{cell}} + \frac{0.059}{6} \log \frac{a_{\text{Al}(s)}^2 \cdot a_{\text{Ni}^{2+}(aq)}^3}{a_{\text{Al}^{3+}(aq)}^2 \cdot a_{\text{Ni}(s)}^3}$$

Now, since Ni and Al both are solids, their activities are equal to unity i.e.  $a_{\text{Ni}} = 1$  and  $a_{\text{Al}} = 1$  for dilute sol<sup>n</sup>, the Nernst equation becomes.

$$E_{\text{cell}} = E^\circ_{\text{cell}} + \frac{0.059}{6} \log \frac{[\text{Ni}^{2+}(aq)]^3}{[\text{Al}^{3+}(aq)]^2}$$

The given cell is composed of two electrodes, which are  $\text{Fe}(s) | \text{Fe}^{2+}(aq, 0.001M)$  [anode] and  $\text{H}^+(aq, 1M) | \text{H}_2(g)$  [cathode] Rxn. taking place at the two electrodes are,

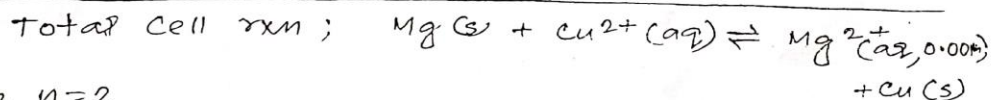
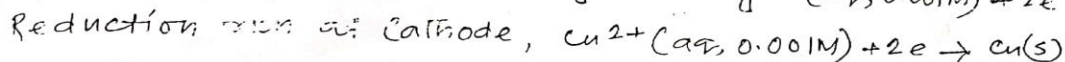
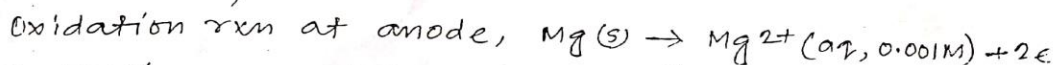


Here  $n=2$

Nernst equation of this rxn can be written as

$$\begin{aligned} E_{\text{cell}} &= E^{\circ}_{\text{cell}} + \frac{0.059}{2} \log \frac{a_{\text{Fe}(s)} \cdot a_{\text{H}^+}^2(aq, 1M)}{a_{\text{Fe}^{2+}}(aq) \cdot a_{\text{H}_2}(g)} \\ &= 0.44 + \frac{0.059}{2} \log \frac{1}{0.001} \\ &= 0.44 + \frac{0.059}{2} \log 10^3 \\ &= 0.528 \text{ V.} \end{aligned}$$

iii) the given cell is composed of two electrodes, which are  $\text{Mg}(s) | \text{Mg}^{2+}(aq, 0.001M)$  [anode] and  $\text{Cu}^{2+}(aq, 0.001M) | \text{Cu}(s)$  [cathode] rxn. taking place at the two electrodes are,



Here  $n=2$

Nernst equation for this cell can be written as, at 298K,

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} + \frac{0.059}{2} \log \frac{a_{\text{Mg}(s)} \cdot a_{\text{Cu}^{2+}}(aq)}{a_{\text{Cu}(s)} \cdot a_{\text{Mg}^{2+}}(aq)}$$

Now since Mg and Cu both are solids, their activities are equal to unity  $a_{\text{Mg}} = 1 = a_{\text{Cu}}$ .

For dil soln. The Nernst equation becomes,

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} + \frac{0.059}{2} \log \frac{[\text{Cu}^{2+}(aq)]}{[\text{Mg}^{2+}(aq)]}$$

$$\begin{aligned} E^{\circ}_{\text{cell}} &= 0.34 - (-2.37) \\ &= 2.71 \end{aligned}$$

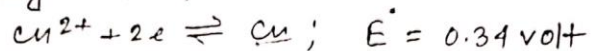
$$\begin{aligned} E_{\text{cell}} &= 2.71 + \frac{0.059}{2} \log \frac{0.001}{0.001} \\ &= 2.71 + \frac{0.059}{2} \log 10^{-1} \\ &= 2.6805 \text{ volt.} \end{aligned}$$

## • Application of standard Redox Potential values.

### i) Prediction of Rxns.

The oxidant of a couple with a more positive standard potential (red) value will in general, oxidise the reductant of a couple with a less positive or negative standard (red) potential value.

Considered the standard electrode potential of the following couples,



The less positive potential of the couple, the more effective the metal is as a reducing agent. Since the  $\text{Cu}^{2+}/\text{Cu}$  couple has a less positive potential Cu will serve as a better reducing agent. Therefore Cu will pass into solution with greater ease than Ag metal.

Therefore if a Cu wire is dipped into an aqueous sol<sup>n</sup> of  $\text{AgNO}_3$ , Cu will pass into solution as  $\text{Cu}^{2+}$  leaving electrons on its surface. These electrons will react with  $\text{Ag}^{+}$  ions which will be reduced to Ag. Effectively there will be a deposit of Ag on the Cu wire and the colourless aqueous solution will assume a blue colour due to  $\text{Cu}^{2+}$  ions. On the contrary when a Ag wire is dipped into  $\text{Cu}^{2+}$  solution there will be no rxn.

Similarly when a Zn rod is dipped into a solution containing  $\text{Cu}^{2+}$  ions, there is a deposit of metallic Cu on the Zn rod ( $E^{\circ}_{\text{Zn}^{2+}/\text{Zn}} = -0.763$ ). On the other hand if a Cu rod is immersed in a solution of  $\text{Ag}^{2+}$  sol<sup>n</sup> no rxn occurs.

### ii) Calculation of Cell potentials.

With the help of  $E^{\circ}$ red values, we can calculate the standard potential of a given galvanic cell. The value of  $E^{\circ}_{\text{cell}}$  is given by -

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{red (right)}} - E^{\circ}_{\text{red (left)}}$$

Let us consider the standard electrode potentials of the following two couples



thus for the cell  $\text{Zn} / \text{Zn}^{2+} \parallel \text{Ag}^{+} / \text{Ag}$ ;

$$\begin{aligned} \text{The cell potential} &= 0.799 - (-0.763) \\ &= 1.562 \text{ volt.} \end{aligned}$$

It is not necessary to consider the no. of electrons involved in the change.

### iii) Determination of equilibrium constant of a redox rxn.

The relation between the standard free energy change and equilibrium constant of a redox rxn given by,

$$\Delta G^{\circ} = -RT \ln K_{\text{eq}} \quad \text{--- (i)}$$

where  $K_{\text{eq}}$  = Equilibrium constant of a redox rxn.

$R$  = molar gas constant

$T$  = Absolute temp

$\Delta G^{\circ}$  = Standard free energy change of the rxn.

$$\text{Again we know, } \Delta G^{\circ} = -nFE^{\circ} \quad \text{--- (ii)}$$

where,  $E^{\circ}$  = Standard EMF of the cell.

$n$  = No. of electrons involved in this rxn.

$F$  = Faraday.

Comparing equ<sup>n</sup> (i) and (ii) we get,

$$nFE^{\circ} = RT \ln K_{\text{eq}}$$

$$\therefore \ln K_{\text{eq}} = \frac{nFE^{\circ}}{RT}$$

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

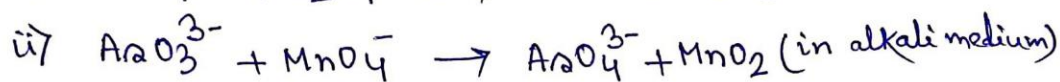
$$\text{or, } \log_{10} K_{\text{eq}} = \frac{nFE^{\circ}}{2.303RT} = \frac{nE^{\circ}}{0.059} \quad \text{at } 298\text{K}$$

$$\text{or, } \boxed{K_{\text{eq}} = 10^{nE^{\circ}/0.059}}$$

Thus knowing  $E^{\circ}$  as well as  $n$  we may calculate the value of equm<sup>n</sup> constant of a redox rxn.

## : Assignment :

1) Balance the following equation by ion electron method:-



2. Predict whether the cell  $\text{Cu}|\text{Cu}^{2+}||\text{Zn}^{2+}|\text{Zn}$  will operate or not?  $E^\circ_{\text{Cu}|\text{Cu}^{2+}} = -0.34\text{V}$

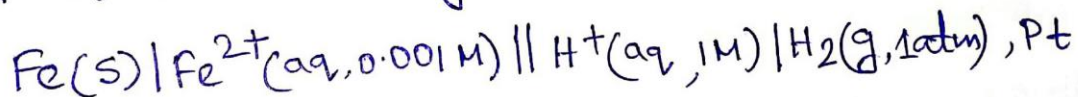
$$E^\circ_{\text{Zn}|\text{Zn}^{2+}} = 0.76\text{V}$$

3. Calculate the equilibrium constant for the overall reaction of the cell  $\text{Sn}^{2+}|\text{Sn}^{4+}||\text{Fe}^{3+}|\text{Fe}^{2+}$

$$\text{Given } E^\circ_{\text{Sn}^{4+}|\text{Sn}^{2+}} = 0.15\text{V}$$

$$E^\circ_{\text{Fe}^{3+}|\text{Fe}^{2+}} = 0.77\text{V}$$

4. Write Nernst equation and calculate EMF of the following cells at 298 K.



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$$\text{given - } E^\circ_{\text{Fe}^{2+}|\text{Fe}} = -0.44\text{V}$$