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REDOX REACTIONS

Modern Concept of Oxidation and Reduction

There are two modern concepts of oxidation and reduction-

(1) Electronic concept

(2) Oxidation number concept

(1) Electronic concept of Oxidation and Reduction

(a) Oxidation

" The process of removal of one or more electrons from any atom, ion or molecule is called oxidation."

For example,

$$Na \rightarrow Na^{+} + e^{-}$$
$$Zn \rightarrow Zn^{2+} + 2e^{-}$$
$$Fe^{2+} \rightarrow Fe^{3+} + e^{-}$$
$$Cl^{-} \rightarrow \frac{1}{2} Cl_2 + e^{-}$$
$$\frac{1}{2} H_2 \rightarrow H^{+} + e^{-}$$

It is clear that, oxidation is a de-electronation process, where positive charge on the species increases or negative charge on the species decreases. This increase or decrease of charge is equal to the number of electrons removed.

(b) Reduction

"The process of gaining of one or more electrons by any atom, ion or molecule is called reduction."

For example,

$$Cl + e^{-} \rightarrow Cl^{-}$$
$$O + 2e^{-} \rightarrow O^{2-}$$
$$Cu^{2+} + 2e^{-} \rightarrow Cu$$
$$Fe^{3+} + e^{-} \rightarrow Fe^{2+}$$
$$\frac{1}{2} H_2 + e^{-} \rightarrow H^{-}$$
$$\frac{1}{2} Cl_2 + e^{-} \rightarrow Cl^{-}$$

It is clear that, reduction is an electronation process, where negative charge on the species increases or positive charge on the species decreases. This increase or decrease of charge is equal to the number of electrons gained.

Oxidation-Reduction Reactions or Redox Reactions

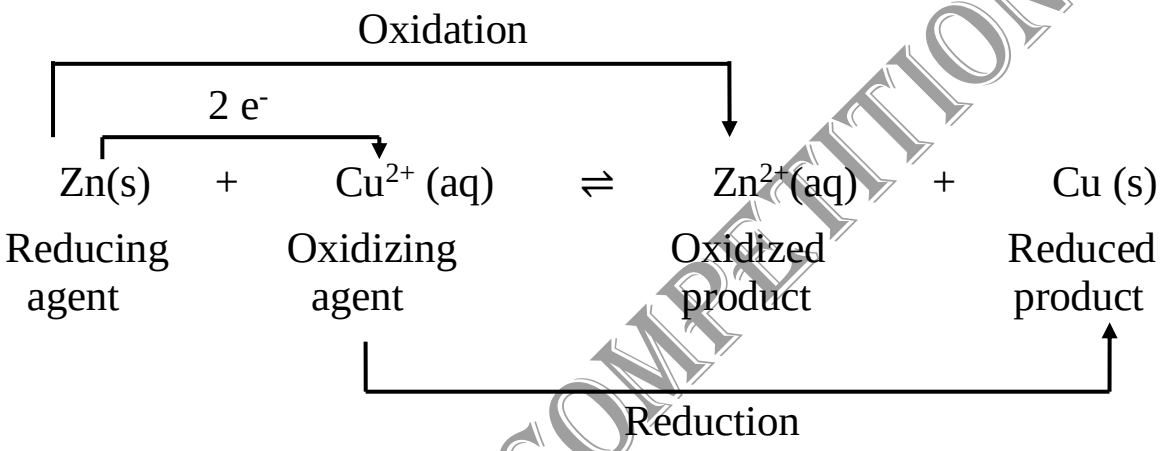
Those reactions in which both oxidation and reduction processes occur simultaneously are called oxidation-reductions reactions or redox reactions."

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In electronic concept oxidation-reductions reactions or redox reactions have been defined in terms of transfer of electron from one reactant to another. By definition,

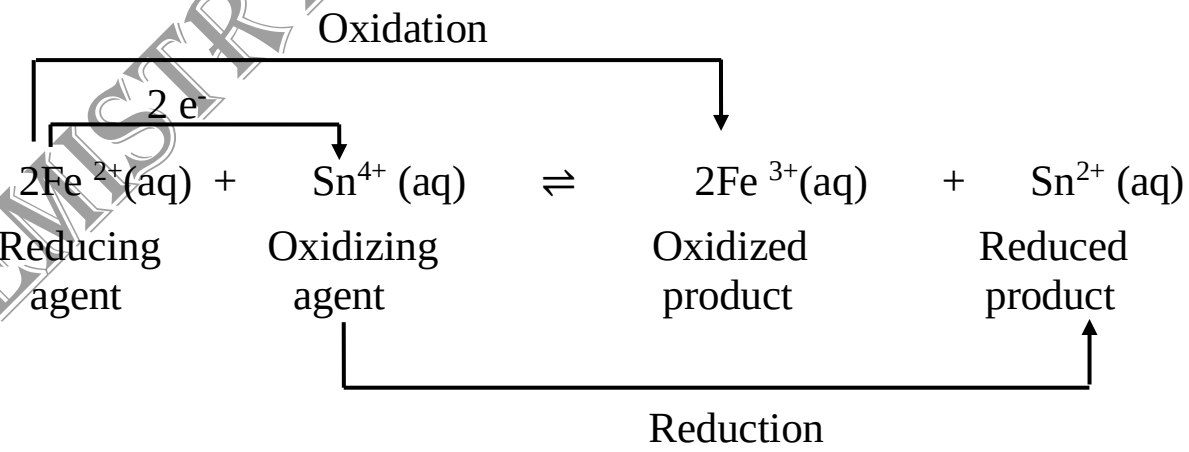
"The reactions in which one or more electrons are transferred from one reactant to another are called oxidation-reductions reactions or redox reactions."

Here, the reactant which looses electrons i.e. which is oxidized causes reduction of another reactant and therefore it is called reducing agent. On the other hand, the reactant which gains electrons i.e. which is reduced causes oxidation of first reactant and therefore it is called oxidizing agent. For example, in following redox reaction-



Here, Zn (s) transfers its two electrons to Cu^{2+} (aq) and gets oxidized into Zn^{2+}(aq). Hence, it acts like reducing agent. On the other hand, Cu^{2+} (aq) gets reduced into Cu(s) by gaining two electrons and acts like oxidizing agent.

Similarly, in following redox reaction-



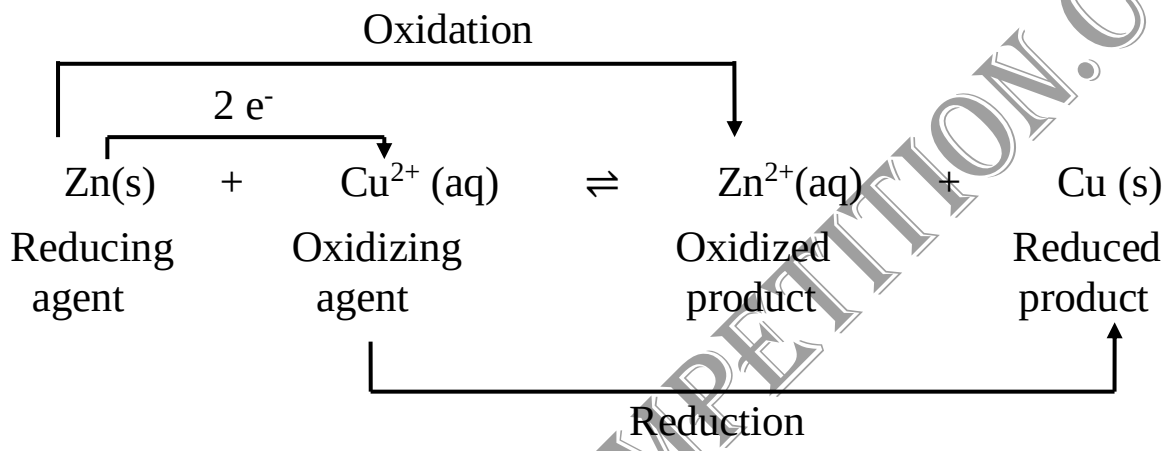
Here, Fe^{2+}(aq) transfers its electron to Sn^{4+} (aq) and gets oxidized into Fe^{3+}(aq). Hence, it acts like reducing agent. On the other hand, Sn^{4+} (aq) gets reduced into Sn^{2+} (aq) by gaining two electrons and acts like oxidizing agent.

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Thus, in redox reactions the electrons are transferred from reducing agent to oxidizing agent.

Oxidation and Reduction Half Reactions

In redox reactions, both oxidation and reduction processes occur simultaneously. Hence, these reactions can be splitted into two half reactions namely oxidation and reduction half reactions. For example, in following redox reaction-



Here, Zn (s) transfers its two electrons to Cu^{2+} (aq) and gets oxidized into Zn^{2+}(aq), while, Cu^{2+} (aq) gets reduced into Cu(s) by gaining two electrons. Thus, oxidation and reduction half reactions involved in this redox reaction will be as under-

$$Zn\ (s) \rightarrow Zn^{2+}\ (aq) + 2e^- \quad \text{(oxidation half reaction)}$$

$$Cu^{2+}\ (aq) + 2e^- \rightarrow Cu\ (s) \quad \text{(reduction half reaction)}$$

$$Zn(s) + Cu^{2+}\ (aq) \rightleftharpoons Zn^{2+}(aq) + Cu\ (s) \quad \text{(Complete redox reaction)}$$

Similarly, other redox reactions can also be splitted into two half reactions, namely oxidation and reduction half reactions.

Types of Redox Reactions

Redox reactions may be of following two types-

(1) Direct redox reactions

(2) Indirect redox reactions

(1) Direct Redox Reactions

Such redox reactions in which both oxidation and reduction processes occur in same vessel are called direct redox reactions.

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For example, if a zinc rod is dipped in aqueous solution of copper sulphate, the rod starts dissolving and a reddish brown precipitate of metallic copper gets deposited on the bottom of the vessel due to following direct redox reaction occurring in the vessel.

$$Zn(s) + CuSO_4(aq) \rightleftharpoons ZnSO_4(aq) + Cu(s)$$

OR

$$Zn(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightleftharpoons Zn^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$$

OR

$$Zn(s) + Cu^{2+}(aq) \rightleftharpoons Zn^{2+}(aq) + Cu(s)$$

This reaction is highly exothermic and therefore the solution gets heated up but the electrical neutrality of the solution remains unaffected.

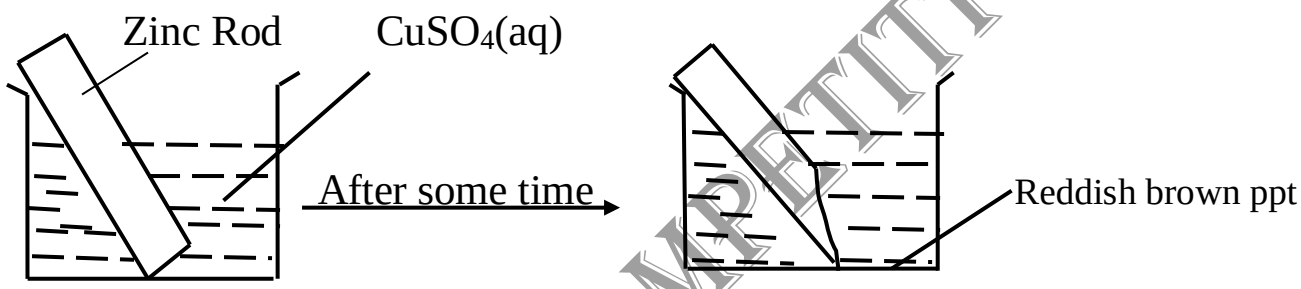


Fig.- Expression of direct redox reaction

Here, reducing and oxidizing agents remain in direct contact of each other. Hence, the transfer of electrons occurs directly from reducing agent to oxidizing agent. Thus, there is no possibility of the flow of free electrons and therefore the electricity cannot be generated. Hence, direct redox reactions cannot form the basis of electrochemical cells.

(2) Indirect Redox Reactions

Such redox reactions in which both oxidation and reduction processes occur in different vessel are called indirect redox reactions.

Here, reducing and oxidizing agents are not in direct contact of each other. Hence, the transfer of electrons cannot occurs directly from reducing agent to oxidizing agent. In fact, in these reactions the electrons are released by reducing agent in one vessel and are transferred in another vessel through conducting wire, where they are gained by oxidizing agent. Thus, there occurs the flow of free electrons and therefore, the electricity is generated.

It is clear that, the indirect redox reactions form the basis of electrochemical cells.

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Electrochemical Cell

It is a device which converts chemical energy into electrical energy. Here, a redox reaction is carried out in indirect manner and decrease in free energy during the reaction is observed as electrical energy.

Let us consider a simple electrochemical cell consisting of zinc and copper electrodes. Here, a zinc rod remains dipped in aqueous solution of zinc sulphate while, a copper rod remains dipped in aqueous solution of copper sulphate. Both of these two rods remain connected together through voltmeter and both the solutions remain chemically connected using salt bridge.

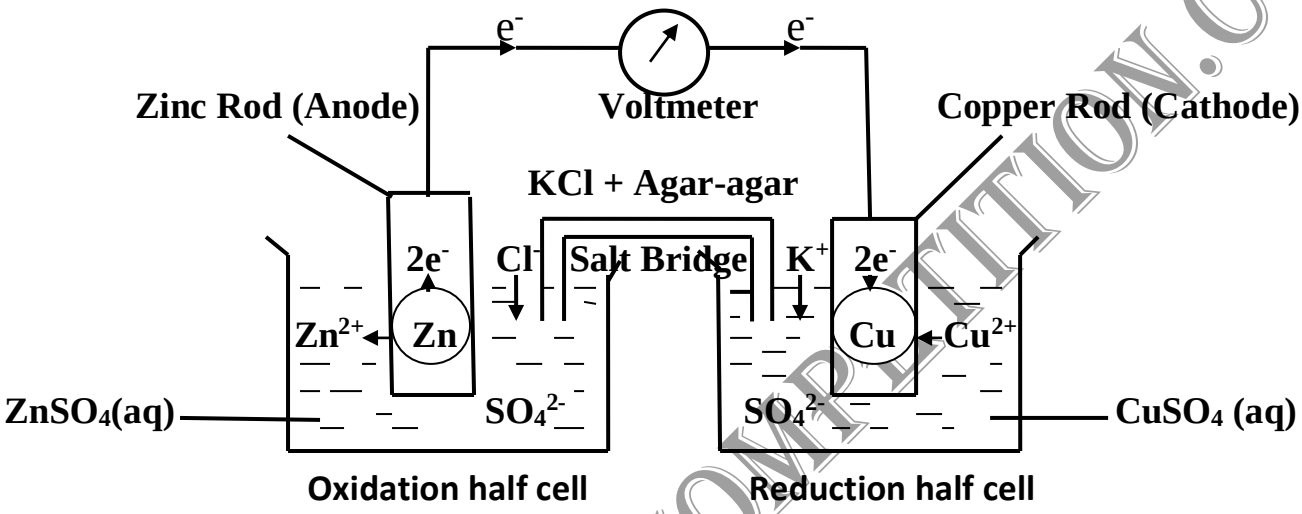


Fig. – Electrochemical Cell

During cell operation, zinc atoms of zinc rod lose their valence electrons on rod and come into solution as $Zn^{2+}(aq)$ ions. These electrons move towards copper rod through conducting wires where, they are taken by $Cu^{2+}(aq)$ ions of solution. Thus, the electrode reactions and over all cell reaction occurring in cell will be as under-

At anode- $Zn\,(s) \rightarrow Zn^{2+}\,(aq) + 2e^-$ (oxidation half reaction)

At cathode $Cu^{2+}\,(aq) + 2e^- \rightarrow Cu\,(s)$ (reduction half reaction)

$$Zn(s) + Cu^{2+}\,(aq) \rightleftharpoons Zn^{2+}(aq) + Cu\,(s)$$

(Complete cell reaction)

Here, oxidation occurs at zinc electrode and therefore, it acts like anode, while reduction occurs at copper electrode and therefore, it acts like cathode. Furthermore, since the electrons are released on anode, it is negatively charged while, they are taken from cathode, it is positively charged. Thus, in electrochemical cells the anode is negatively charged while, the cathode is positively charged.

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Here, due to flow of electrons from the anode to the cathode electricity is generated, which flows in the direction opposite to the direction of flow of electrons i.e. from the cathode to the anode.

Salt Bridge and its Function

It is an inverted U-tube consisting of glass and remains filled with a semi-solid paste of an inert electrolyte such as KCl, KNO_3, $NaNO_3$, NH_4Cl etc with agar-agar. Its one end remains dipped in anodic solution, while another end in cathodic solution. Both of these two ends remain provided with cotton plug to provide mechanical support to the paste.

Functions of Salt Bridge

Salt bridge performs following two important functions in an electrochemical cell-

(1) It complete the circuit by providing chemical contact to the anodic and cathodic solutions, so that, the flow of electrons occurs continuously in the cell and therefore, electricity is generated continuously.

(2) It maintains electrical neutrality of anodic and cathodic solutions by furnishing appropriate number of anions and cations in these solutions, so that flow of electrons occurs continuously in the cell and therefore, electricity is generated continuously.

If salt bridge is not used, then after some time, anodic solution will become positively charged and cathodic solution will become negatively charged. Hence, the flow of electrons will obstructed and generation of electricity will get stopped.

Symbolic Representation of Electrochemical Cell

Following are international conventions for the symbolic representation of an electrochemical cell-

(1) Anode is always represented on left hand side while, the cathode is represented on right hand side.

(2) During representation of anode metal is represented first and it is followed by metal ion. Both of these two are separated either by semi colon or by single vertical line.

$$M(s) \mid M^{n+}(aq)$$

Anode

For example, zinc electrode acting like anode can be represented as under-

$$Zn(s) \mid Zn^{2+}(aq)$$

Anode

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(3) During representation of cathode metal ion is represented first and it is followed by metal. Both of these two are separated either by semi colon or by single vertical line.

$$M^{n+}(aq) \mid M(s)$$

Cathode

For example, copper electrode acting like cathode can be represented as under-

$$Cu^{2+}(aq) \mid Cu(s)$$

Cathode

(4) Both anode and cathode are separated by two parallel vertical lines which represents the salt bridge.

(5) Some time, molar concentration of electrolytes is also represented in parentheses after the symbol of metal ion.

For example, zinc-copper cell involving following cell reaction-

$$Zn(s) + Cu^{2+}(aq) \rightleftharpoons Zn^{2+}(aq) + Cu(s)$$

can be symbolically represented as under-

$$Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$$

Anode Salt bridge Cathode

Q.- Give the symbolic representation of an electrochemical cell involving following cell reaction-

$$Cu(s) + 2Ag^{+}(aq) \rightleftharpoons Cu^{2+}(aq) + 2Ag(s)$$

Ans.- Here, copper electrode is acting like anode, while, silver electrode is acting like cathode. Hence, the electrochemical cell involving given cell reaction can be symbolically represented as under-

$$Cu(s) \mid Cu^{2+}(aq) \parallel Ag^{+}(aq) \mid Ag(s)$$

Anode Salt bridge Cathode

Q.- Give the electrode reactions and the complete cell reaction occurring in following electrochemical cell-

$$Cu(s) \mid Cu^{2+}(aq) \parallel Ag^{+}(aq) \mid Ag(s)$$

Ans.- In given electrochemical cell, copper electrode is acting like anode, while silver electrode is acting like cathode. Hence, the electrode reactions and the complete cell reaction occurring in this electrochemical cell can be represented as under-

At anode- $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^{-}$ (oxidation half reaction)

At cathode $2Ag^{+}(aq) + 2e^{-} \rightarrow 2Ag(s)$ (reduction half reaction)

$$Cu(s) + 2Ag^{+}(aq) \rightleftharpoons Cu^{2+}(aq) + 2Ag(s)$$

(complete cell reaction)

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Electrode Potential

If a metal rod is dipped in aqueous solution of its salt, then after some times the rod becomes either positively or negatively charged with respect to the solution. Thus, a charge difference (potential difference) gets developed between metal rod and solution. This potential difference is called electrode potential of the respective electrode.

For example, if a zinc rod is dipped in aqueous solution of $ZnSO_4$, then after some times the rod becomes negatively charged with respect to the solution i.e. a negative electrode potential gets developed on zinc electrode.

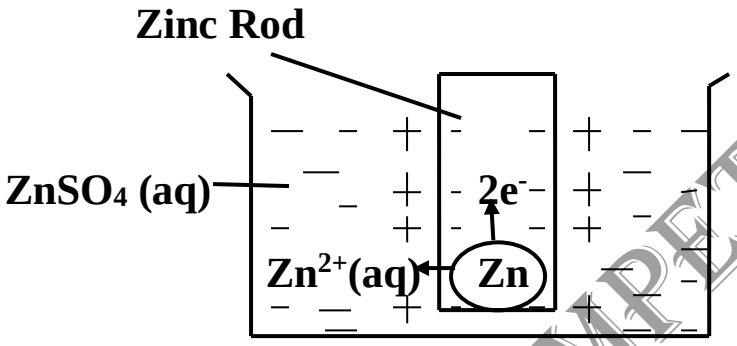


Fig.- Development of negative electrode potential on zinc electrode

Similarly, if a copper rod is dipped in aqueous solution of $CuSO_4$, then after some times the rod becomes positively charged with respect to the solution i.e. a positive electrode potential gets developed on copper electrode.

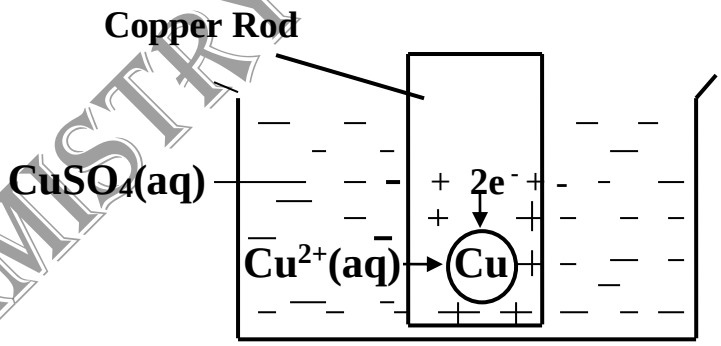


Fig.- Development of positive electrode potential on copper electrode

In fact, if a metal rod is dipped in aqueous solution of its salt, following two important cases may arise-

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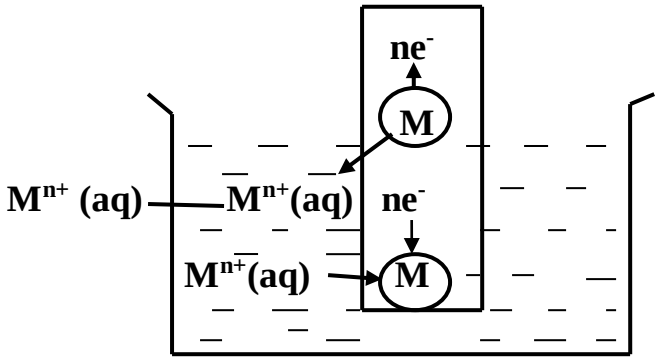


Fig. – Development of electrode electrode potential

(1) Metal atoms of metal rod lose their electrons on rod and come into solution as metal ions.

$$M\ (s) \rightleftharpoons M^{n+} + ne^-$$

This reaction is reversible, where at equilibrium the metal rod is negatively charged due to presence of excess of electrons. Thus, a negative electrode potential gets developed on metal rod. For example, zinc electrode.

(2) Metal ions of solution take electrons from rod and get reduced into metal atoms. These metal atoms either get deposited on metal rod or get settled on the bottom.

$$M^{n+} + ne^- \rightleftharpoons M\ (s)$$

This reaction is reversible, where at equilibrium the metal rod is positively charged due to deficiency of electrons. Thus, a positive electrode potential gets developed on metal rod. For example, copper electrode.

Types of Electrode Potential

Electrode potential may be of following two types-

(1) Oxidation Potential

The tendency of an electrodes to get oxidized i.e. to give electrons is called is called oxidation electrode potential of the electrode. For example,

$$M\ (s) \rightleftharpoons M^{n+} + ne^-$$
$$Zn\ (s) \rightleftharpoons Zn^{2+} + 2e^-$$
$$\tfrac{1}{2}\ H_2\ (g) \rightleftharpoons H^+ + e^-$$

Oxidation electrode potential is the measure of capacity of the electrode to get oxidized. The larger the value of oxidation electrode potential, more is capacity of the electrode to get oxidized.

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(2) Reduction Potential

The tendency of an electrodes to get reduced i.e. to take electrons is called is called reduction electrode potential of the electrode. For example,

$$M^{n+} + ne^- \rightleftharpoons M\ (s)$$
$$Cu^{2+} + 2e^- \rightleftharpoons Cu\ (s)$$
$$H^+ + e^- \rightleftharpoons \tfrac{1}{2} H_2\ (g)$$

Reduction electrode potential is the measure of capacity of the electrode to get reduced. The larger the value of reduction electrode potential, more is capacity of the electrode to get reduced.

All the electrodes have both oxidation and reduction electrode potentials. The numerical values of these two electrode potentials are also same for an electrode but they have opposite sign. Thus,

$$E_{oxidation} = - E_{reduction}$$

For example, standard oxidation electrode potential of zinc electrode is +0.76V, hence its standard reduction electrode potential will be –0.76V. Similarly, standard oxidation electrode potential of copper electrode is -0.34 V, hence its standard reduction electrode potential will be +0.34 V.

Following Europian convention, in India, standard reduction potential is taken as standard electrode potential of the electrode.

Standard Hydrogen Electrode (SHE)

It is a primary standard reference electrode, which is used to measure relative standard electrode potential of different electrode. It contains a platinum wire fused in a thin glass tube and is remains fitted with a platinum foil coated with platinum black. Thin glass tube remains jacketed with a hard glass jacket having opening for passing hydrogen gas. The platinum foil remains dipped in 1M HCl solution in such a way that its two-third part remains dipped in solution, while one-third remains exposed to hydrogen gas being constantly bubbled in solution at 298 K temperature and 1 atm pressure.

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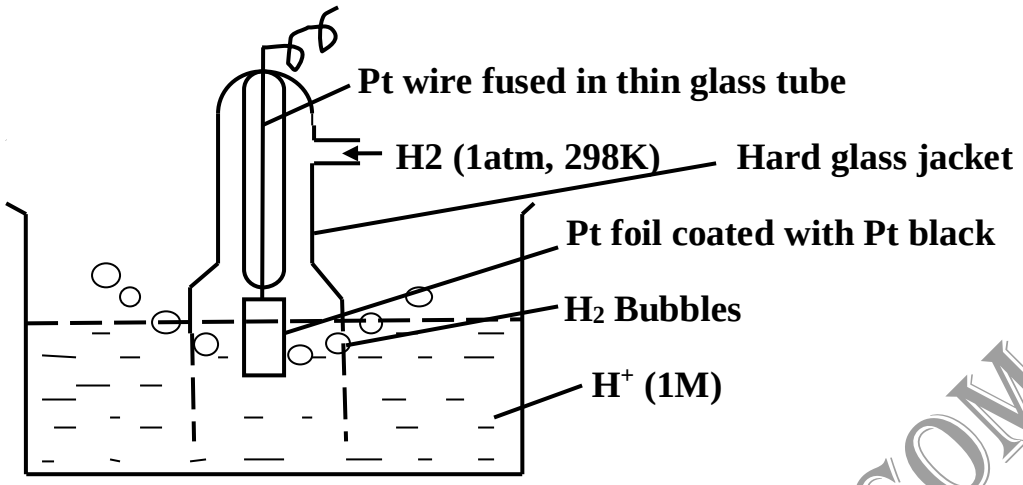


Fig.- Standard Hydrogen Electrode (SHE)

SHE is a primary standard reversible electrode, which can act both like anode and cathode.

(a) When SHE acts like anode, then oxidation occurs on this electrode.

$$\frac{1}{2} H_2 \rightleftharpoons H^+ + e^- \quad \text{(oxidation)}$$

Here, this electrode is represented as under-

$$Pt, H_2(1atm) \mid H^+(1M)$$

Anode

(b) When SHE acts like cathode, then reduction occurs on this electrode.

$$H^+ + e^- \rightleftharpoons \frac{1}{2} H_2 \quad \text{(reduction)}$$

Here, the electrode is represented as under-

$$H^+(1M) \mid H_2(1atm), Pt$$

Cathode

The standard electrode potential of SHE has been arbitrarily fixed to be $\pm$ 0.00 V and that of other electrodes is determined with respect to this electrode.

Determination of Standard electrode potential

For experimental determination of Standard electrode potential of an electrode, the electrode in standard conditions (298 K, 1 atm) is connected to SHE to make an electrochemical cell and its standard cell potential is determined. Now knowing the standard cell potential (E^0_{cell}) and standard electrode potential of one of the electrodes, that of another electrode can be calculated using following formula-

$$E^0_{cell} = E^0_{cathode} - E^0_{anode}$$

For example, let we have to determine the standard electrode potential of zinc electrode. Here, zinc electrode in standard conditions is connected with SHE through a voltmeter to make an electrochemical cell. Here, the direction of flow of electrons is from zinc electrode to SHE, which indicates that, zinc electrode is

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acting like anode and SHE is acting like cathode. Hence, electrode reactions and overall cell reaction occurring in cell will be as follows-

At anode- $Zn\,(s) \rightarrow Zn^{2+}\,(aq) + 2e^-$ (oxidation half reaction)
At cathode $2H^+\,(aq) + 2e^- \rightarrow H_2\,(g)$ (reduction half reaction)

$$Zn(s) + 2H^+\,(aq) \rightleftharpoons Zn^{2+}(aq) + H_2\,(g)$$

(complete cell reaction)

The cell representation of this electrochemical cell will be as under-

$$Zn(s) \mid Zn^{2+}(aq) \parallel 2H^+\,(aq) \mid H_2\,(g), Pt$$

Here, from voltmeter, standard cell potential of the cell,

$E^0_{cell} = 0.76\ V$

$\therefore \quad E^0_{cell} = E^0_{H^+/H2} - E^0_{Zn^{2+}/Zn}$

or $\quad 0.76\ V = 0.00V - E^0_{Zn^{2+}/Zn}$

or $\quad E^0_{Zn^{2+}/Zn} = -\,o.76\ V$

Thus, standard electrode potential of zinc electrode is - o.76 V.

Similarly, let we have to determine the standard electrode potential of copper electrode. Here, copper electrode in standard conditions is connected with SHE through a voltmeter to make an electrochemical cell. Here, the direction of flow of electrons is from SHE to copper electrode, which indicates that, SHE is acting like anode and copper electrode is acting like cathode. Hence, electrode reactions and overall cell reaction occurring in cell will be as under-

At anode- $H_2\,(g) \rightarrow 2H^+\,(aq) + 2e^-$ (oxidation half reaction)
At cathode $Cu^{2+}\,(aq) + 2e^- \rightarrow Cu(s)$ (reduction half reaction)

$$H_2\,(g) + Cu^{2+}\,(aq) \rightleftharpoons 2H^+\,(aq) + Cu(s)$$

(complete cell reaction)

The cell representation of this electrochemical cell will be as under-

$$Pt, H_2\,(g) \mid H^+(aq) \parallel Cu^{2+}\,(aq) \mid Cu(s)$$

Here, from voltmeter, standard cell potential of the cell,

$E^0_{cell} = 0.34\ V$

$\therefore \quad E^0_{cell} = E^0_{Cu^{2+}/Cu} - E^0_{SHE}$

or $\quad 0.34\ V = E^0_{Cu^{2+}/Cu} - 0$

or $\quad E^0_{Cu^{2+}/Cu} = +\,o.34\ V$

Thus, standard electrode potential of copper electrode is + o.34 V.

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Nernst Equation for Single Electrode Potential

Nernst equation gives the effect of temperature and concentration on electrode potential of an electrode. Let, an electrode involves following electrode reaction of reduction-

$$M^{n+}(aq) + ne^- \rightarrow M(s)$$

Following Nernst equation, the electrode potential of this electrode in a condition other than standard condition will be as follows-

$$E_{M^{n+}/M} = E^0_{M^{n+}/M} - \frac{RT}{nF} \ln \frac{[PRODUCTS]}{[REACTANTS]}$$

$$\therefore \quad E_{M^{n+}/M} = E^0_{M^{n+}/M} - \frac{RT}{nF} \ln \frac{[M(s)]}{[M^{n+}(aq)]}$$

Where, $E^0_{M^{n+}/M}$ = standard electrode potential of the electrode

$E_{M^{n+}/M}$ = electrode potential of the electrode in a condition other than Standard condition

n = Number of electrons involved in electrode reaction

R = Gas constant ($8.314 JK^{-1}mol^{-1}$)

F = Faradey's constant (96500 Coulombs)

T = Absolute temperature

By conventions, [M(s)] = 1

$$\therefore \quad E_{M^{n+}/M} = E^0_{M^{n+}/M} - \frac{RT}{nF} \ln \frac{1}{[M^{n+}(aq)]} \quad \text{-------} \quad (1)$$

$$\text{or} \quad E_{M^{n+}/M} = E^0_{M^{n+}/M} + \frac{RT}{nF} \ln [M^{n+}(aq)] \quad \text{-------} \quad (2)$$

Changing the base of logarithms from e to 10, we get-

$$E_{M^{n+}/M} = E^0_{M^{n+}/M} - \frac{2.303RT}{nF} \log \frac{1}{[M^{n+}(aq)]} \quad \text{-------} \quad (3)$$

$$\text{or} \quad E_{M^{n+}/M} = E^0_{M^{n+}/M} + \frac{2.303RT}{nF} \log [M^{n+}(aq)] \quad \text{-------} \quad (4)$$

At 298K temperature, we get following simplified form of Nernst equation-

$$E_{M^{n+}/M} = E^0_{M^{n+}/M} - \frac{0.059}{n} \log \frac{1}{[M^{n+}(aq)]} \quad \text{-------} \quad (5)$$

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$$\text{or} \quad E_{M^{n+}/M} = E^0_{M^{n+}/M} + \frac{0.059}{n} \log [M^{n+}(aq)] \quad \text{------- (6)}$$

All these equations are Nernst equations for single electrode potential of different electrodes.

Nernst Equation for Cell Potential

Nernst equation gives the effect of temperature and concentration on electrode potential of the electrodes and therefore, it also gives the effect of temperature and concentration on cell potential of the cells. Let us consider following electrochemical cell-

$$Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$$

The cell reaction occurring in this cell is as under-

$$Zn(s) + Cu^{2+}(aq) \rightleftharpoons Zn^{2+}(aq) + Cu(s)$$

Following Nernst equation, the cell potential of this electrochemical cell in a condition other than standard condition will be as under-

$$E_{cell} = E^0_{cell} - \frac{RT}{nF} \ln \frac{[PRODUCTS]}{[REACTANTS]}$$

$$\therefore \quad E_{cell} = E^0_{cell} - \frac{RT}{2F} \ln \frac{[Cu(s)][Zn^{2+}(aq)]}{[Cu^{2+}(aq)][Zn(s)]}$$

Where, E^0_{cell} = standard cell potential of the cell

E_{cell} = cell potential of the cell in a condition other than standard condition

n = Number of electrons involved in electrode reaction

R = Gas constant ($8.314 JK^{-1}mol^{-1}$)

F = Faradey's constant (96500 Coulombs)

T = Absolute temperature

By conventions, $[Zn(s)] = [Cu(s)] = 1$

$$\therefore \quad E_{cell} = E^0_{cell} - \frac{RT}{2F} \ln \frac{[Zn^{2+}(aq)]}{[Cu^{2+}(aq)]} \quad \text{------- (1)}$$

$$\text{or} \quad E_{cell} = E^0_{cell} + \frac{RT}{2F} \ln \frac{[Cu^{2+}(aq)]}{[Zn^{2+}(aq)]} \quad \text{------- (2)}$$

Changing the base of logarithms from e to 10, we get-

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$$E_{cell} = E^0_{cell} - \frac{2.303RT}{2F} \log \frac{[Zn^{2+}(aq)]}{[Cu^{2+}(aq)]} \quad \text{------- (3)}$$

$$\text{or} \quad E_{cell} = E^0_{cell} + \frac{2.303RT}{2F} \log \frac{[Cu^{2+}(aq)]}{[Zn^{2+}(aq)]} \quad \text{------- (4)}$$

At 298K temperature, we get following simplified form of Nernst equation-

$$E_{cell} = E^0_{cell} - \frac{0.059}{2} \log \frac{[Zn^{2+}(aq)]}{[Cu^{2+}(aq)]} \quad \text{------- (5)}$$

$$\text{or} \quad E_{cell} = E^0_{cell} + \frac{0.059}{2} \log \frac{[Cu^{2+}(aq)]}{[Zn^{2+}(aq)]} \quad \text{------- (6)}$$

All these equations are Nernst equations for cell potential of zinc-copper cell.

Similarly, for following electrochemical cell-

$$Cu(s) \mid Cu^{2+}(aq) \parallel Ag^{+}(aq) \mid Ag(s)$$

The cell reaction occurring in this cell is as under-

$$Cu(s) + 2Ag^{+}(aq) \rightleftharpoons Cu^{2+}(aq) + 2Ag(s)$$

Following Nernst equation, the cell potential of this electrochemical cell in a condition other than standard condition will be as under-

$$E_{cell} = E^0_{cell} - \frac{RT}{nF} \ln \frac{[PRODUCTS]}{[REACTANTS]}$$

$$\therefore E_{cell} = E^0_{cell} - \frac{RT}{2F} \ln \frac{[Cu^{2+}(aq)][Ag(s)]^2}{[Cu(s)][Ag^{+}(aq)]^2}$$

By conventions, $[cu(s)] = [Ag(s)] = 1$

$$\therefore \quad E_{cell} = E^0_{cell} - \frac{RT}{2F} \ln \frac{[Cu^{2+}(aq)]}{[Ag^{+}(aq)]^2} \quad \text{------- (1)}$$

Changing the base of logarithms from e to 10, we get-

$$E_{cell} = E^0_{cell} - \frac{2.303RT}{2F} \log \frac{[Cu^{2+}(aq)]}{[Ag^{+}(aq)]^2} \quad \text{------- (2)}$$

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At 298K temperature, we get following simplified form of Nernst equation-

$$E_{cell} = E^0_{cell} - \frac{0.059}{2} \log \frac{[Cu^{2+}(aq)]}{[Ag^{+}(aq)]^2} \quad \text{------- (3)}$$

All these equations are Nernst equations for cell potential of copper-silver cell.

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