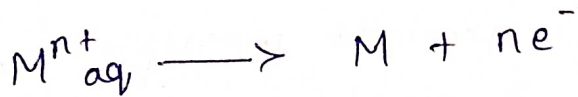


NERNST EQUATION :

Consider the following electrode reaction.



The emf of a reversible cell can be measured by free energy change (ΔG) of a reaction takes place in the reversible cell. If the reaction involves the transfer of 'n' electrons, 'F' Faradays of electricity will flow and E is the emf of the cell, then the total electrical energy produced by the cell is given by following equation

$$-\Delta G = nFE \quad (1)$$

where $-\Delta G$ is decrease in free energy change.

In other way,

$$\Delta G^{\circ} = -nFE^{\circ} \quad \text{or} \quad (-)\Delta G^{\circ} = nFE^{\circ} \quad (2)$$

where $-\Delta G^{\circ}$ is Standard free energy change and E° is the Standard emf of a cell.

For a reversible reaction, the interrelationship of free energy change and equilibrium Constant (K) is given by

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

$$\Delta G = (-)RT \ln K + RT \ln \frac{[\text{product}]}{[\text{reactant}]}$$

But $\Delta G^{\circ} = (-)RT \ln K \quad (2)$

Sub

$$\Delta G = \Delta G^{\circ} + RT \ln \frac{[\text{product}]}{[\text{reactant}]}$$

where n is the number of electrons

F is the Faraday = 96,500 Coulombs of electricity

E° is the Standard potential

Sub the equation

$$(-)nFE = (-)nFE^\circ + RT \ln \frac{[\text{product}]}{[\text{reactant}]}$$

Rearranging the above equation.

$$E_{\text{red}} = E^\circ_{\text{red}} - RT \ln \frac{[\text{product}]}{[\text{reactant}]}$$

$$E_{\text{red}} = E^\circ_{\text{red}} - \frac{RT}{nF} \ln \frac{[M]}{[M^{n+}]}$$

$[M] = 1$ for the solid metal.

$$E_{\text{red}} = E^\circ_{\text{red}} - \frac{RT}{nF} \ln \frac{1}{[M^{n+}]}$$

The above equation may be written as,

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

where $R = 8.314 \text{ J/K/mole}$; $T = 298^\circ\text{C}$; $F = 96500 \text{ Coulombs}$.

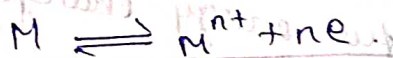
Sub the value of R , T , and F .

$$E_{\text{red}} = E^\circ_{\text{red}} + \frac{0.0591}{n} \ln [M^{n+}]$$

The above equation is known as Nernst Equation.

For reduction reaction $E_{\text{red}} = E^\circ_{\text{red}} + \frac{0.0591}{n} \ln [M^{n+}]$

For oxidation reaction,



$$E_{\text{oxi}} = E^\circ_{\text{oxi}} - \frac{0.0591}{n} \ln [M^{n+}]$$

Application of Nernst Equation:

- * Corrosion tendency of metals can be predicted.
- * It is used to calculate the EMF of a cell.
- * pH of a solution can be calculated by measuring the emf.
- * Nernst Equation is used to calculate electrode potential of unknown metal.
- * Concentration of the reactant can be calculated using the electrode potential.