

Electrochemistry B.Sc. (II) sub.  
D.B. College (Jaynagar)  
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### 9. NERNST EQUATION:

With the help of Nernst equation, we can calculate the electrode potential of electrode or EMF of cell

$$E_{\text{cell}} = E^{\circ} - \frac{RT}{nF} \log_e \frac{[\text{Product}]}{[\text{Reactant}]}$$

Where -  $E^{\circ}$  = Standard electrode potential

$R$  = gas constant

$T$  = temperature (in K)

$F$  = Faraday (96500 Coulomb  $\text{mol}^{-1}$ )

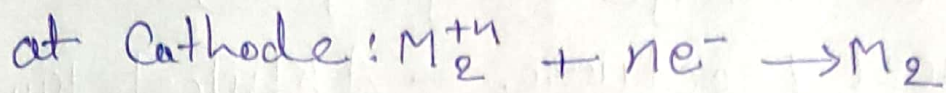
$n$  = no. of  $e^{-}$  gained or lost in balanced equation.

$$\text{or, } E_{\text{cell}} = E^{\circ} - \frac{0.591}{n} \log_{10} \frac{[\text{Product}]}{[\text{Reactant}]} = E^{\circ} - \frac{0.0591}{n} \log Q$$

Let, in the cell:

at Anode:  $M_1 \rightarrow M_1^{+n} + ne^{-}$   
for this reaction -

$$E_{\text{oxi}} = E_{\text{oxi}}^{\circ} - \frac{0.0591}{n} \log_{10} \frac{[M_1^{+n}]}{[M_1]} \quad \text{--- (1)}$$



for this reaction —

$$E_{\text{red.}} = E_{\text{red.}}^{\circ} - \frac{0.0591}{n} \log_{10} \frac{[M_2]}{[M_2^{+n}]} \quad \text{--- (2)}$$

Note: Concentration of Solid taken as unity. --- (2)

$$\text{So } [M_1] = [M_2] = 1$$

We know that EMF of cell is

$$EMF = E_{\text{oxi}} + E_{\text{red}}$$

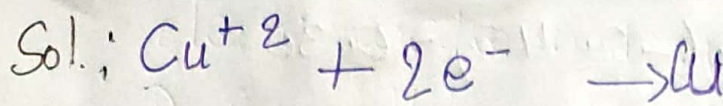
(Anode)                  (Cathode)

By adding equation (1) and (2) we get:

$$\begin{aligned} EMF &= E_{\text{oxi}}^{\circ} - \frac{0.0591}{n} \log [M_1^{+n}] + E_{\text{red.}}^{\circ} - \frac{0.0591}{n} \log \frac{1}{[M_2^{+n}]} \\ &= (E_{\text{oxi}}^{\circ} + E_{\text{red.}}^{\circ}) - \frac{0.0591}{n} [\log [M_1^{+n}] - \log [M_2^{+n}]] \end{aligned}$$

$$EMF = E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log \frac{[M_1^{+n}]}{[M_2^{+n}]}$$

Ex: The 0.1M copper sulphate solution in which copper electrode is dipped at 25°C. Calculate the electrode potential of copper electrode [Given  $E^{\circ} \text{Cu}^{+2} / \text{Cu} = 0.34\text{V}$ ]



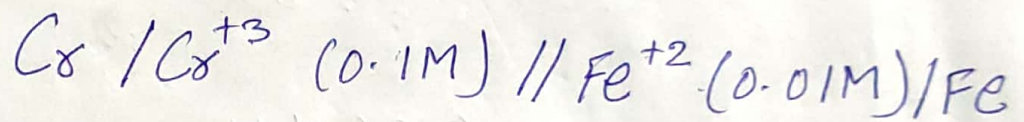
$$E_{\text{red}} = E_{\text{red}}^{\circ} - \frac{0.0591}{n} \log \frac{[\text{Product}]}{[\text{Reactant}]}$$

here  $n = 2$

So  $E = 0.34 - \frac{0.0591}{2} \cdot \log_{10}$

$$= 0.34 - 0.03 = 0.31 \text{ Volts}$$

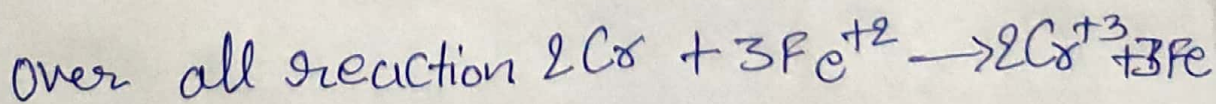
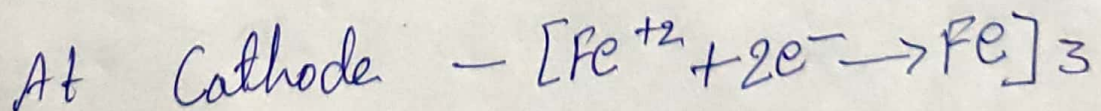
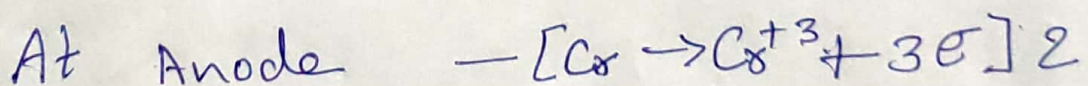
Ex: The EMF of Cell



(Given  $E_{\text{Cr}^{+3} / \text{Cr}} = -0.75\text{V}$

$E_{\text{Fe}^{+2} / \text{Fe}} = 0.45\text{V}$ )

Sol.: Half Cell reactions are:



$E_{\text{cell}} = \text{Oxidation Pot.} + \text{Reduction Pot.}$

$$= 0.75 + (-0.45) = 0.30$$

$$E_{\text{cell}} = E^{\circ} - \frac{0.0591}{n} \log \frac{[\text{Product}]}{[\text{Reactant}]}$$

$$= 0.30 - \frac{0.0591}{6} \log \frac{[\text{Cr}^{+3}]^2}{[\text{Fe}^{+2}]^3}$$

$$= 0.30 - \frac{0.0591}{6} \log \frac{[0.1]^2}{[0.01]^3}$$

$$= 0.30 - \frac{0.24}{6}$$

$$= 0.26 \text{ Volt.}$$