

# DETERMINATION OF K USING NERNST EQUATION



- When the reaction is at **equilibrium**, no net reaction occurs and **no net transfer of electron occurs**.

$$\Rightarrow E_{\text{cell}} = 0, \text{ and } Q = K$$

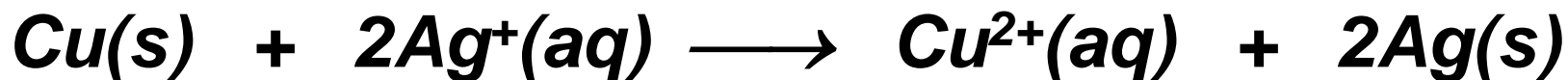
- Nernst Equation :

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0592}{n} \log Q$$

$$E^{\circ}_{\text{cell}} = \frac{0.0592}{n} \log K$$

## Example

Calculate the equilibrium constant (K) for the following reaction.



**Solution :**

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$$

$$= E^{\circ}_{\text{Ag}^+/\text{Ag}} - E^{\circ}_{\text{Cu}^{2+}/\text{Cu}}$$

$$= +0.80 - 0.34$$

$$= +0.46 \text{ V}$$

At equilibrium,  $E_{\text{cell}} = 0$ ,  $Q = K$

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0592}{n} \log Q$$

$$0 = E^{\circ}_{\text{cell}} - \frac{0.0592}{2} \log K$$

$$E^{\circ}_{\text{cell}} = \frac{0.0592}{2} \log K$$

$$0.46 = \frac{0.0592}{2} \log K$$

$$\log K = 15.54$$

$$K = 3.467 \times 10^{15}$$