Electrode potential and cell emf

Electrode Potential

- **Origin of Electrode Potential:**
  - Arises due to the tendency of a metal to lose or gain electrons when in contact with its own ion solution.
  - Metal atoms tend to dissolve as cations, leaving behind excess electrons, creating a potential difference.
  - The equilibrium between oxidation and reduction reactions defines the electrode potential.
- **Nernst Equation:**
  - Used to calculate the electrode potential of a half-cell under non-standard conditions.
  - Formula:

    $$E = E^\circ - \frac{RT}{nF}\ln Q$$

    Where:
    - E = Electrode potential at given conditions
    - E° = Standard electrode potential
    - R = Universal gas constant (8.314 J/mol·K)
    - T = Temperature in Kelvin
    - n = Number of electrons involved
    - F = Faraday's constant (96485 C/mol)
    - Q = Reaction quotient
- **Example:**
  - Consider a Zn/Zn^{2+} half-cell at 25°C with $[Zn^{2+}] = 0.01M$.
  - Standard electrode potential: $E^\circ_{\mathrm{Zn^{2+}/Zn}} = -0.76V$
  - Using Nernst equation:

$$E = -0.76 - \frac{0.0591}{2}\log(0.01)$$

$$E = -0.76 + 0.02955 \times 2$$

$$E = -0.70V$$

- **Formula for Cell EMF:**

$$E_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

- **Problem:**
  Calculate the cell potential for a Daniell cell at 25°C when $[Zn^{2+}] = 0.001M$ and $[Cu^{2+}] = 0.1M$.
  Given:
  - $E^\circ_{\mathrm{Zn^{2+}/Zn}} = -0.76V$
  - $E^\circ_{\mathrm{Cu^{2+}/Cu}} = 0.34V$

Solution:

1. Calculate standard EMF:

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

$$= 0.34 - (-0.76) = 1.10V$$

2. Apply Nernst equation:

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{2} \log \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

$$= 1.10 - \frac{0.0591}{2} \log \frac{0.001}{0.1}$$

$$= 1.10 - \frac{0.0591}{2} \log(10^{-2})$$

$$= 1.10 - \frac{0.0591}{2} \times (-2)$$

$$= 1.10 + 0.0591 = 1.16V$$

Final Answer: $E_{\text{cell}} = 1.16V$