CONCEPT: NERNST EQUATION

The ***Nernst Equation*** reveals the quantitative connection between the concentrations of compounds and cell potential.

$$E = E^o - \frac{RT}{nF}\ln\frac{A_B^b}{A_A^a}$$

E^o = Standard Cell Potential

R = Gas Constant = ______________________

n = Number of electrons transferred

F = Faraday's constant = __________________

A = Activity

______ represents the cell potential under non-standard conditions, while ______ represents it under standard conditions.

At 25 °C, $$\frac{RT}{F} = \frac{\left(8.314\frac{J}{K\cdot mol}\right)(298.15\ K)}{\left(9.649\times10^4\frac{C}{mol}\right)} = 0.0257\frac{J}{C} = 0.0257\ V$$

The *Nernst Equation* then becomes,

$$E_{Cell} = E^o_{Cell} - \frac{0.0257\ V}{n}\ln Q$$

By multiplying ln by 2.303 we can obtain the log function.

$$E_{Cell} = E^o_{Cell} - \frac{0.05916\ V}{n}\log Q$$

The cell potential calculated from Nernst equation is the maximum potential at the instant the cell circuit is connected. As the cell discharges and current flows, the electrolyte concentrations will change, *Q* increases and E_{cell} deceases.

- Over time the reaction will reach equilibrium at then Q = K and cell potential will equal zero.

$$E_{Cell} = E^o_{Cell} - \left(\frac{RT}{nF}\right)\ln K = 0$$

$$\Delta G = \Delta G^o - RT\ln K = 0$$

PRACTICE: NERNST EQUATION CALCULATIONS 1

EXAMPLE 1: Determine the cell potentials of the following concentration cells:

Ag | Ag^+ (*aq*, 0.0010 M) || Cu^{2+} (aq, 1.0 M) | Cu $E^{o}_{-} = 0.799$ V $E^{o}_{+} = 0.339$ V

EXAMPLE 2: Consider the following electrochemical cell for the question:

Pt | KBr(aq, 0.01 M), Br_2(l) || $FeBr_2$(aq, 1.0M) | Fe $E^{o}_{-} = 1.078$ V $E^{o}_{+} = -0.440$ V

Determine the spontaneity and cell potential based on the given cell notation.

PRACTICE: NERNST EQUATION CALCULATIONS 2

EXAMPLE 1: Consider a standard voltaic cell based on the reaction:

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

Which of the following actions would change the emf of the cell?

a) Increasing the pH at the cathode

b) Lowering the pH at the cathode

c) Increasing $[Sn^{2+}]$ at the anode

d) Increasing the hydrogen gas pressure at the cathode

e) All of the above changes will alter the cell potential

EXAMPLE 2: Consider the following half cell reaction at T = 25. °C

$$Ni^{2+}(aq) + 2\ e^{-} \longrightarrow Ni(s) \qquad E° = -\ 0.260\ V$$

What will be the value for E°, the half cell potential for standard conditions, for the reaction

$$2\ Ni(s) \longrightarrow 2\ Ni^{2+}\ (aq) + 4\ e^{-}$$

a) + 0.52 V

b) + 0.26 V

c) – 0.26 V

d) – 0.52 V