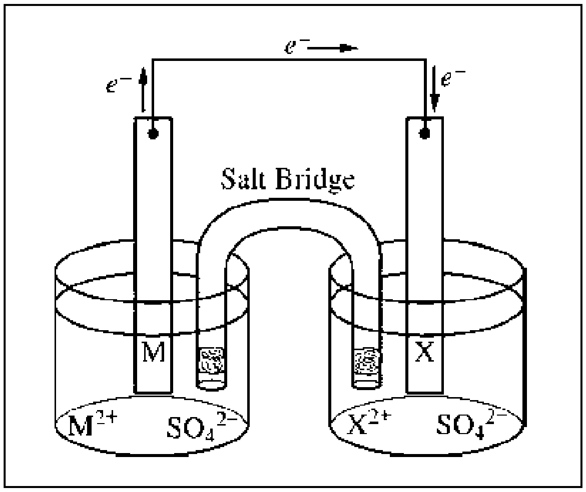
AP CHEMISTRY

TOPIC 11: ELECTROCHEMISTRY, REVIEW, PART III

Day 132:

1. The diagram below shows the experimental setup for a typical electrochemical cell that contains two standard half-cells. The cell operates according to the reaction represented by the following equation.

$$Zn\ (s)\ +\ Ni^{+2}\ (aq)\ \longrightarrow\ Ni\ (s)\ +\ Zn^{+2}\ (aq)$$



(a) Identify M and M^{+2} in the diagram and specify the initial concentration for M^{+2} in solution.

Answers:

$$\boldsymbol{M = Zn,\ \ M^{+2} = Zn^{+2},\ \ Zn^{+2} = [\ 1.0\ ] = 1\ \text{molar concentration}}$$

(b) Indicate which of the metal electrodes is the cathode. Write the balanced equation for the reaction that occurs in the half-cell containing the cathode.

Answers:

Cathode is where reduction takes place…

$$\boldsymbol{Ni^{+2}\ +\ 2\ e^{-1}\ \longrightarrow\ Ni}$$

(c) What would be the effect on the cell voltage if the concentration of Zn^{+2} was reduced to 0.100 *M* in the half-cell containing the Zn electrode? (do not use a calculator !)

Answers:

$$E_{cell} = E^0_{cell} - \frac{0.0592}{n}\ \log\ Q$$

$$Q = \frac{\left[Zn^{+2}\right]}{\left[Ni^{+2}\right]}\ \ ;\ \ \log\ \frac{[0.10]}{[1.0]} = -\ \#$$

Since the value of log Q results in a negative value, the cell voltage will be INCREASED.

(d) Describe what would happen to the cell voltage if the salt bridge was removed. Explain.

Answers:

The voltage would go (drop) to zero. This happens because the salt bridge is needed to allow charge balance to occur in the solutions the electrodes are immersed in. In the absence of the salt bridge, ions cannot flow to balance the buildup of cations in the anode compartment and the buildup of anions in the cathode compartment. Also, if the buildup in charge is allowed to occur, then the flow of electrons will be affected (stop) - the two half reactions must be occurring for the voltage to exist.

2. Answer the following questions that relate to electrochemical reactions.

(a) Under standard conditions at 25^0C, Zn (*s*) reacts with Co^{+2} (*aq*) to produce Co (*s*).

(i) Write the balanced equation for the oxidation half reaction.

Answers:

$$Zn \longrightarrow Zn^{+2} + 2e^{-1}$$

(ii) Write the balanced net-ionic equation for the overall reaction.

Answers:

$$Co^{+2} + Zn \longrightarrow Zn^{+2} + Co$$

(iii) Calculate the standard potential, E^0, for the overall reaction at 25^0C.

Answers:

$$Co^{+2} + 2e^{-1} \longrightarrow Co \qquad E^0 = -\,0.28\ V$$

$$Zn \longrightarrow Zn^{+2} + 2e^{-1} \qquad E^0 = +\,0.76\ V$$

$$E^0 = +\,0.76\ V - 0.28\ V = +\,0.48\ V$$

(b) At 25^0C, H_2O_2 decomposes according to the following equation.

$$2\ H_2O_2\ (aq) \longrightarrow 2\ H_2O\ (l) + O_2\ (g) \quad E^0 = 0.55\ V$$

(i) Determine the value of the standard free energy change, ΔG , for the reaction at 25^0C.

Answers:

$$2\ \overset{+1}{H}_2\overset{-1}{O}_2 \rightarrow 2\ \overset{+1}{H}_2\overset{-2}{O} + \overset{0}{O_2}$$

$$2e^{-1} + 2\,O^{-1} \longrightarrow O^{-2}$$

$$2\,O^{-1} \longrightarrow O_2 + 2e^{-1}$$

$$\boxed{\begin{array}{c} n = 2 \\ \text{2 electrons are exchanged} \end{array}}$$

$$\Delta G = -n\Im E^0$$

$$\Delta G = -\left(2\ mol\ e^-\right)\left(\frac{96500\ J}{mol\ e^- \cdot V}\right)\left(0.55\ V\right) = -1.06\times10^5\ J$$

(ii) Determine the value of the equilibrium constant, K, for the reaction at 25^0C.

Answers: You know everything but "K"

$$\Delta G = -RT \ \ln(K)$$

$$-1.06\times10^5 \ J = -\left(\frac{8.31 \ J}{mol\cdot K}\right)(298 \ K) \ \ln(K)$$

$$\frac{-1.06\times10^5 \ J}{-\left(\frac{8.31 \ J}{mol\cdot K}\right)(298 \ K)} = \ln(K) = \frac{1.06\times10^5 \ J(mol\cdot K)}{(8.31 \ J)(298 \ K)} = 42.9$$

$$K = e^{42.865} = 4.13\times10^{18}$$

(iii) The standard reduction potential, E^0, for the half-reaction $O_2\ (g) + 4\ H^{+1}\ (aq) + 4\ e^{-1} \longrightarrow 2\ H_2O\ (l)$ has a value of 1.23 V. Using this information in addition to the information given above, determine the value of the standard reduction potential, E^0, for the half reaction below.

$$O_2\ (g) + 2\ H^{+1}\ (aq) + 2\ e^{-1} \longrightarrow H_2O_2\ (aq)$$

Answers:

$$\boldsymbol{O_2 + 4\ H^{+1} + 4\ e^{-1} \longrightarrow 2\ H_2O \qquad E^0 = 1.23\ V}$$

$$\boldsymbol{2\ H_2O + O_2 \longrightarrow 2\ H_2O_2 \qquad E^0 = -\ 0.55\ V}$$

$$\boldsymbol{2\ O_2 + 4\ H^{+1} + 4\ e^{-1} \longrightarrow 2\ H_2O_2 \qquad E^0 = +\ 0.68\ V}$$

$$\boldsymbol{O_2 + 2\ H^{+1} + 2\ e^{-1} \longrightarrow H_2O_2 \qquad E^0 = +\ 0.68\ V}$$

(c) In the electrolytic cell. Cu (s) is produced by the electrolysis of $CuSO_4$ (aq). Calculate the maximum mass of Cu (s) that can be deposited by a direct current of 100. amperes passed through 5.00 L of 2.00 M $CuSO_4$ (aq) for a period of 1.00 hour.

$$\frac{1 \ hour}{} \times \frac{60 \ \min}{1 \ hr} \times \frac{60 \ \sec}{1 \ \min} = 3600 \ \sec$$

$$I = \frac{q}{t} \ ; \ q = It = (100.\ amp)(3600 \ \sec) = 360000 \ amp\cdot\sec$$

$$\frac{360000 \ amp\cdot\sec}{} \times \frac{1 \ mol \ e^-}{96500 \ amp\cdot\sec} \times \frac{1 \ mol \ Cu^{+2}}{2 \ mol \ e^-} \times \frac{63.55 \ g}{1 \ mol \ Cu^{+2}} = 119 \ g$$

Recall, we DO NOT care about how many total moles of Cu^{+2} ions we have (5.00 L of 2.00 M) – only how much SOLID copper that can be produced !!!