Electrochemistry Review #2

1. Balance the following redox reactions:

a. $Zn(s) + VO^{+2}(aq) \rightarrow Zn^{+2}(aq) + V^{+3}(aq)$

acid

ox: $Zn \rightarrow Zn^{+2} + 2e^-$

red: $(e^- + 2H^+ + VO^{+2} \rightarrow V^{+3} + H_2O)\,2$

$Zn + \cancel{2e^-} + 4H^+ + 2VO^{+2} \rightarrow 2V^{+3} + 2H_2O + Zn^{+2} + \cancel{2e^-}$

$Zn + 4H^+ + 2VO^{+2} \rightarrow 2V^{+3} + 2H_2O + Zn^{+3}$

b. $Zn(s) + VO_3^-(aq) \rightarrow V^{+2}(aq) + Zn^{+2}(aq)$

base

ox: $(Zn \rightarrow Zn^{+2} + 2e^-)\,3$

red: $(3e^- + 6H^+ + VO_3^- \rightarrow V^{+2} + 3H_2O)\,2$

$\cancel{6e^-} + 12H^+ + 2VO_3^- + 3Zn \rightarrow 3Zn^{+2} + \cancel{6e^-} + 2V^{+2} + 6H_2O$

$12OH^- + 12H^+ + 2VO_3^- + 3Zn \rightarrow 3Zn^{+2} + 2V^{+2} + \cancel{6H_2O} + 12OH^-$

$6H_2O + 2VO_3^- + 3Zn \rightarrow 3Zn^{+2} + 2V^{+2} + 12OH^-$

2. Calculate the voltage delivered by a voltaic cell using the following reaction if all dissolved species are 2.5×10^{-2} M.

$$Sn(s) + 2Ag^+(aq) \rightarrow Sn^{+2}(aq) + 2Ag(s)$$

$Sn \rightarrow Sn^{+2} + 2e^-$ .14

$(Ag^{+1} + e^- \rightarrow Ag)\,2$ .8

.94

$E_{cell} = .94 - \frac{.0591}{2} \log\left(\frac{2.5\times10^{-2}}{(2.5\times10^{-2})^2}\right)$

$E_{cell} = .893\ V$

3. A voltaic cell is constructed using the reaction below:

$$Mg(s) + 2H^+(aq) \rightarrow Mg^{+2}(aq) + H_2(g)$$

a. Write equations for the oxidation and reduction half reactions.

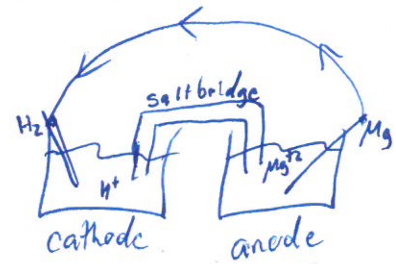
ox: $Mg \rightarrow Mg^{+2} + 2e^-$

red: $2e^- + 2H^+ \rightarrow H_2$

b. Determine which half-reaction occurs at the anode and which is at the cathode.

c. Diagram the cell indicating electron flow.

d. Determine the cell potential if all conditions are standard.



$E_{cell} = 2.37 + 0 \rightarrow 2.37\ V$

4. One half-cell in a voltaic cell is constructed with a silver wire dipped into a solution of $AgNO_3$ of unknown concentration. The other half-cell consists of a zinc electrode in a 1 M solution of $Zn(NO_3)_2$. A voltage of 1.48 V is measured for this cell. Use this information to determine the concentration of the $Ag(NO_3)$ solution.

$Ag^+ + e^- \rightarrow Ag$ .8

~~$Zn^{+2} + 2e^- \rightarrow Zn$ -.76~~

$Zn \rightarrow Zn^{+2} + 2e^-$ +.76

$2Ag^+ + Zn \rightarrow 2Ag + Zn^{+2}$

$1.48 = 1.56 - \frac{.0591}{2} \log\left[\frac{1}{x^2}\right]$

$\log\left(\frac{1}{x^2}\right) = 2.7073$

$\frac{1}{x^2} = 509.65$

$x = .044\ M\ Ag^{+1}$

5. One half-cell in a voltaic cell is constructed with a copper wire dipped into a $4.8x10^{-3}$M solution of $Cu(NO_3)_2$. The other half-cell consists of a zinc electrode in a 0.40 M solution of $Zn(NO_3)_2$. Calculate the cell potential.

$Cu^{+2} + 2e^- \rightarrow Cu \quad .34$

$\cancel{Zn^{+2} + 2e^- \rightarrow Zn \quad -.76}$

$Zn \rightarrow Zn^{+2} + 2e^- \quad +.76$

$E^\circ_{cell} : 1.1V$

$E_{cell} = 1.1 - \frac{.0591}{2} \log\left(\frac{.4}{4.8\times10^{-3}}\right)$

$E_{cell} = 1.04 \text{ v}$

6. Consider an electrochemical cell based on the half reactions Al^{+3} (aq) + $3e^-$ → Al(s) and Cu^{+2}(aq) + $2e^-$ → Cu(s).

-1.66 $\quad$ $.34$

a. Diagram the cell and label each of the components (anode, cathode, and salt bridge).

b. Write a balanced equation for the overall cell reaction.

$3Cu^{+2} + 2Al \rightarrow 2Al^{+3} + 3Cu$

c. What is the standard cell potential?

$1.66 + .34 = 2V$

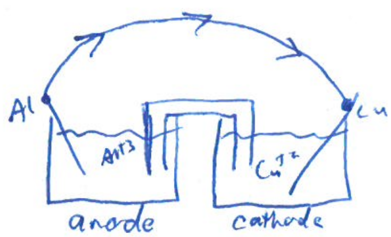
d. In which direction do the electrons flow in the external circuit?

$Al \rightarrow Cu$

e. Assume that the salt bridge contains KNO_3. In which direction do the K^+(aq) ions move? In what direction do the NO_3^-(aq) ions move?

$K^+ \rightarrow$ Cu side $\quad NO_3^- \rightarrow$ Al side

f. If the concentration of Cu^{+2} is reduced to $3.5x10^{-3}$ M, and Ni^{+2} remains 1.0 M, what is the value of Ecell?



f.) $E_{cell} = 2 - \frac{.0591}{6} \log\left(\frac{1^2}{(3.5\times10^{-3})^3}\right)$

$E_{cell} = 1.93 \text{ v}$

7. What mass of solid nickel will be deposited on the cathode from a solution of Ni_{+2}(aq) using a current of 0.150 A for 12.2 minutes. Remember, 1 Amp = 1 coulomb/second; 1 mole of electrons carries a charge of 96,485 Coulombs.

$\frac{.15\,C}{1\,sec} \Big| \frac{12.2\,min}{} \Big| \frac{60\,sec}{1\,min} = 109.8\,C \Big| \frac{1\,mol\,e^-}{96{,}485\,C} \Big| \frac{1\,mol\,Ni}{2\,mol\,e^-} \Big| \frac{58.7\,g\,Ni}{1\,mol\,Ni} = .033\text{ g Ni}$

8. Assume the specifications of a Ni-Cd voltaic cell include delivery of 0.25 A of current for 1.00 h. What is the minimum mass of cadmium that must be used to make the anode in this cell? The cadmium will be placed in a Cd_{+2}(aq) solution. Remember, 1 Amp = 1 coulomb/second; 1 mole of electrons carries a charge of 96,485 Coulombs.

$\frac{.25\,C}{1\,sec} \Big| \frac{3600\,s}{} = 900\,C \Big| \frac{1\,mol\,e^-}{96{,}485\,C} \Big| \frac{1\,mol\,Cd}{2\,mol\,e^-} \Big| \frac{112.4\,g\,Cd}{1\,mol\,Cd} = .524\text{ g Cd}$