AP Chemistry: Electrochemistry

AP Chemistry Electrochemistry WS 0809.doc

Name: Key Date: ____________ Per: ____

1. Zinc metal reacts with nitric acid, HNO_3, to produce a number of products, depending on how dilute the acid solution is. In a concentrated solution, zinc reduces nitrate ion to ammonium ion; zinc is oxidized to zinc ion, Zn^{2+}. Write the net ionic equation for this reaction.

$Zn \rightarrow Zn^{2+} + 2e^-$ $\quad 8e^- + 10H^+ + NO_3^- \rightarrow NH_4^+ + 3H_2O$

$4Zn \rightarrow 4Zn^{2+} + 8e^-$

$[4Zn + 10H^+ + NO_3^- \rightarrow 4Zn^{2+} + NH_4^+ + 3H_2O]$

2. Permanganate ion oxidizes sulfite ion in basic solution according to the following equation:

$$MnO_4^-{}_{(aq)} + SO_3^{2-}{}_{(aq)} \rightarrow MnO_{2(s)} + SO_4^{2-}{}_{(aq)}$$

Use the half-reaction method to complete and balance this equation.

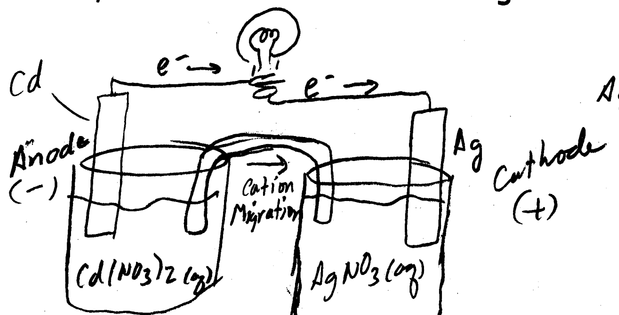
$3e^- + 4H^+ + MnO_4^- \rightarrow MnO_2 + 2H_2O$ $\quad H_2O + SO_3^{2-} \rightarrow SO_4^{2-} + 2H^+ + 2e^-$

$6e^- + 8H^+ + 2MnO_4^- \rightarrow 2MnO_2 + 4H_2O$ $\quad 3H_2O + 3SO_3^{2-} \rightarrow 3SO_4^{2-} + 6H^+ + 6e^-$

$2H^+ + 2MnO_4^- + 3H_2O + 3SO_3^{2-} \rightarrow 2MnO_2 + 4H_2O + 3SO_4^{2-} + 6H^+$ $\quad [H_2O + 2MnO_4^- + 3SO_3^{2-} \rightarrow 2MnO_2 + 3SO_4^{2-} + 2OH^-]$

3. A voltaic cell is constructed from a half-cell in which a cadmium rod dips into a solution of cadmium nitrate, $Cd(NO_3)_2$, and another half-cell in which a silver rod dips into a solution of silver nitrate, $AgNO_3$. The two half-cells are connected by a salt bridge. Silver ions are reduced during the operation of the voltaic cell. Draw a sketch of the cell. Label the anode and cathode, showing the corresponding half-reactions at these electrodes. Indicate the electron flow in the external circuit (with a lightbulb), the signs of the electrodes, and the direction of cation migration in the half-cells.

$Cd_{(s)} \rightarrow Cd^{2+} + 2e^-$



$Ag^+ + e^- \rightarrow Ag_{(s)}$

4. Write the condensed cell notation in which the electrode reactions are:

$2H^+{}_{(aq)} + 2e^- \rightarrow H_{2(g)}$

$Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2e^-$

$Zn \,|\, Zn^{2+}{}_{(aq)} \,||\, H^+{}_{(aq)} \,|\, H_{2(g)} \,|\, Pt$

5. Write the cell reaction for the voltaic cells:

a. $Tl_{(s)} \,|\, Tl^+{}_{(aq)} \,||\, Sn^{2+}{}_{(aq)} \,|\, Sn_{(s)}$

Ox: $Tl_{(s)} \rightarrow Tl^+ + e^-$ $\quad$ Red: $Sn^{2+} + 2e^- \rightarrow Sn_{(s)}$

b. $Zn_{(s)} \,|\, Zn^{2+}{}_{(aq)} \,||\, Fe^{3+}{}_{(aq)}, Fe^{2+}{}_{(aq)} \,|\, Pt$

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

Red: $Fe^{3+}{}_{(aq)} + e^- \rightarrow Fe^{2+}{}_{(aq)}$

6. The cell potential of a particular voltaic cell with the cell reaction:

$Hg_2^{2+}{}_{(aq)} + H_{2(g)} \leftrightarrow 2Hg_{(l)} + 2H^+{}_{(aq)}$

is 0.650V. Calculate the maximum electrical work of this cell when 0.500g of H_2 is consumed.

$\Delta G^o = W_{max} = -nFE$

$= (2)(96,480)(0.650V) = -125,424\ J$

$.500g\ H_2 / 2.016\ g/mol = .248\ mol\ H_2$ $\quad -125,424J(.248\ mol) = [-3.11\times10^4\ J]$

7. a. Order the following oxidizing agents by increasing strength under standard-state conditions: $Cl_{2(g)}$, $H_2O_{2(aq)}$, $Fe^{3+}{}_{(aq)}$

$Fe^{3+}{}_{(aq)} < Cl_{2(g)} < H_2O_{2(aq)}$ (increasing half-reduction potentials)

b. Order the following reducing agents by increasing strength under standard-state conditions: $H_{2(g)}$, $Al_{(s)}$, $Cu_{(s)}$

$Cu < H_2 < Al$ increasingly negative reduction potential in their ion states

8. Consider the reaction:

$Zn^{2+}{}_{(aq)} + 2Fe^{2+}{}_{(aq)} \rightarrow Zn_{(s)} + 2Fe^{3+}{}_{(aq)}$

Does the reaction go spontaneously in the direction indicated, under standard conditions?

No. Zinc's reduction potential is lower than iron's.

$Fe^{3+} \rightarrow Fe^{2+}\ (0.77V)$ $\quad Zn^{2+} \rightarrow Zn\ (-0.76V)$

9. Using standard electrode potentials, calculate the standard free-energy change at 25.0°C for the reaction:

$Zn_{(s)} + 2Ag^+{}_{(aq)} \rightarrow Zn^{2+}{}_{(aq)} + 2Ag_{(s)}$

What happens to the free-energy if the coefficients of the reaction are doubled? What about the cell potential?

$E^o_{cell} = (0.80 + 0.76) = 1.56V$

$\Delta G^o = -nFE^o_{cell} = -(2)(96,480)(1.56) = -301017.6\ J \approx [-301\ kJ]$

Free energy doubles; Cell potential remains the same

10. a. What is the cell potential of the following voltaic cell at 25.0°C? $E^o_{cell} = 0.34 + 0.76 = 1.10V$

$Zn_{(s)}\ |\ Zn^{2+}\ (1.00\times10^{-5}M)\ ||\ Cu^{2+}\ (0.100M)\ |\ Cu_{(s)}$

$E_{cell} = E^o_{cell} - \left(\frac{RT}{nF}\right)\ln Q = E^o_{cell} - \left(\frac{0.025693V}{n}\right)\ln Q = 1.10V - \left(\frac{.025693}{2}\right)\ln\left(\frac{1.00\times10^{-5}}{0.100}\right)$

b. What is the equilibrium constant, K_{eq}, for this reaction under standard conditions? $= 1.218 = [1.22V]$

$E^o_{cell} = \frac{RT}{nF}\ln K_{eq}$ $\quad 1.10 = \left(\frac{0.025693}{2}\right)\ln K_{eq}$ $\quad 85.6 = \ln K_{eq}$

$K_{eq} = e^{85.6} = [1.54\times10^{37}]$

11. When an aqueous solution of copper(II)sulfate, $CuSO_4$, is electrolyzed, copper metal is deposited:

$Cu^{2+}{}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

(The other electrode gives oxygen: $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$). If a constant current is passed for 5.00h and 404mg of copper metal was deposited, what was the current?

$\frac{.404g}{63.55\ g/mol} = .0064\ mol\ Cu \left(\frac{2\ mol\ e^-}{1\ mol\ Cu}\right) = .0127\ mol\ e^- \times \left|\frac{96,480\ C}{mol\ e^-}\right| = 1226.685C$

$1226.685C / 18000\ s = .0681\ A = [68.1\ mA]$

12. Write a net ionic equation for the expected reaction when the electrolysis of $NiSO_{4(aq)}$ is conducted using an inert platinum anode and a nickel cathode.

$-0.257V\quad Ni^{2+} + 2e^- \rightarrow Ni$

$+1.229V\quad O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$

$-0.828V\quad 2H_2O + 2e^- \rightarrow H_{2(g)} + 2OH^-{}_{(aq)}$

Nickel gets reduced & the water is oxidized

Since the Pt is the anode it will not be oxidized and the Ni cathode will not be reduced further. Sulfate requires acid to be reduced

$[2Ni^{2+}{}_{(aq)} + 2H_2O_{(l)} \rightarrow 2Ni_{(s)} + O_{2(g)} + 4H^+{}_{(aq)}]$