1. Write the expression for the solubility product constant of $Al(OH)_3$
a) in terms of molarities

$K_{sp} = [Al^{3+}][OH^-]^3$

b) in terms of solubility

$K_{sp} = s\,(3s)^3 = 27s^4$

2. Solubility of $Al(OH)_3$ in pure water is _____greater /_______**lower** than the solubility in NaOH solution.

3. An electrochemical cell in which a non-spontaneous reaction is driven by an external source of direct current is called **electrolytic cell**

4. The oxidation number of Cr in $K_2Cr_2O_7$ is **+6**

5. In the reaction $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$ the reducing agent is **H_2**

6. Write denotation of gas chlorine electrode

Pt ¦ $Cl_2(g)$ ¦ $Cl^-(aq)$

7. In the expression for standard cell potential ν represents **number of moles of electrons**

8. In order to measure the standard potential of electrode one sets up the cell in which the standard electrode is **an anode** and the electrode of interest is **a cathode**

II. Problems

1. At 633 ^oC the equilibrium constant K_c for the dissociation of ammonia to its elements is 6.56×10^{-3} .

$$2\,NH_3(g) == N_2(g) + 3\,H_2(g)$$

a) calculate K_p

$K_p = 36.21$

b) ammonia is placed in a 2 L flask where it generates 456 mm Hg of pressure. When equilibrium has been established, what is the total pressure in the flask?

$p_{tot} = 1.14$ atm

2. K_c at 2000 K for the formation of NO(g) is 4×10^{-4}.

$$N_2(g) + O_2(g) == 2NO(g)$$

a) If analysis shows that the concentrations of N_2 and O_2 are both 0.25 M and that of NO is 0.0042 M, is the system at equilibrium?

NO

b) If the system is not at equilibrium, in which direction does the reaction proceed?

Rxn goes →

c) When the system is at equilibrium, what are the equilibrium concentrations?

$[N_2] = [O_2] = 0.2496$ **$[NO] = 0.005$**

2. The solubility product constant for $AgIO_3$ is 1.0×10^{-8}. If 0.10 g of solid $AgIO_3$ is added to 100 mL of 0.0.20 M KIO_3, what are the concentrations K^+, IO_3^-, and Ag^+ at equilibrium?

$[K^+] = 4.673 \times 10^{-3}$ M **$[Ag^+] = 2.14 \times 10^{-6}$ M** **$[IO_3^-] = 4.675 \times 10^{-3}$ M**

4. Balance the following unbalanced equation:

$$\mathbf{3}Cu(s) + \mathbf{2}NO_3^-(aq) + \mathbf{8}H_3O^+(aq) \dashrightarrow \mathbf{3}Cu^{2+}(aq) + \mathbf{2}NO(g) + \mathbf{12}H_2O(l)$$

5. Consider the following half-reactions:

$Cl_2(g) + 2e \rightarrow 2Cl^-(aq)$ $E^o = +1.36$ V
$I_2(s) + 2e \rightarrow 2I^-(aq)$ $E^o = +0.535$ V
$Pb^{2+}(aq) + 2e \rightarrow Pb(s)$ $E^o = -0.126$ V
$V^{2+}(aq) + 2e \rightarrow V(s)$ $E^o = -1.18$ V

a) which is the weakest oxidizing agent on the list? **V^{2+}**
b) Which is the strongest reducing agent? **V(s)**
c) Will Pb(s) reduce V^{2+}(aq) to V(s)? **NO**
d) Will I_2(s) oxidize Cl^-(aq) to Cl_2(g)? **NO**
e) Name the elementsor ions that can be reduced by Pb(s) **Cl_2(g) and I_2(s)**

6. Write the cell reaction and electrode half-reactions for the following cell:

Pt(s) | Cl_2(g) | HCl(aq) || K_2CrO_4(aq) | Ag_2CrO_4(s) | Ag(s)

$2Cl^-(aq) + 2Ag^+(aq) \rightarrow Cl_2(g) + Ag(s)$

7. Use the standard potentials of the couples Au^+/Au (+1.69 V), Au^{3+}/Au (+1.40 V), and Fe^{3+}/Fe^{2+} (+0.77 V) to calculate E^o and the equilibrium constant for the reaction:

$$2Fe^{2+}(aq) + Au^{3+}(aq) == 2Fe^{3+}(aq) + Au^+(aq)$$

$E^o_{cell} = 0.49$ V **$K = 3.8 \times 10^{16}$**