

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

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

- b) in terms of solubility

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

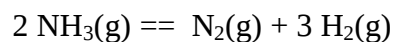
2. Solubility of Al(OH)_3 in pure water is _____greater / _____**lower** than the solubility in NaOH solution.
3. An electrochemical cell in which a non-spontaneous reaction is driven by an external source of direct current is called **electrolytic cell**
4. The oxidation number of Cr in $\text{K}_2\text{Cr}_2\text{O}_7$ is **+6**
5. In the reaction $\text{CuO(s)} + \text{H}_2\text{(g)} \rightarrow \text{Cu(s)} + \text{H}_2\text{O(g)}$ the reducing agent is **H_2**
6. Write denotation of gas chlorine electrode



7. In the expression for standard cell potential v 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°C the equilibrium constant K_c for the dissociation of ammonia to its elements is 6.56×10^{-3} .



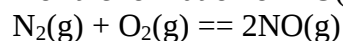
- a) calculate K_p

$$K_p = 36.21$$

- 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?

$$p_{\text{tot}} = 1.14 \text{ atm}$$

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



- 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 $\rightarrow$

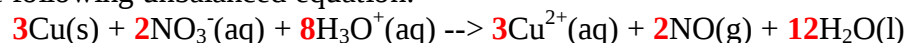
- c) When the system is at equilibrium, what are the equilibrium concentrations?

$[\text{N}_2] = [\text{O}_2] = 0.2496$ $[\text{NO}] = 0.005$

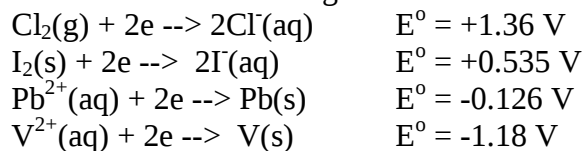
2. The solubility product constant for AgIO_3 is 1.0×10^{-8} . If 0.10 g of solid AgIO_3 is added to 100 mL of 0.020 M KIO_3 , what are the concentrations K^+ , IO_3^- , and Ag^+ at equilibrium?

$[\text{K}^+] = 4.673 \times 10^{-3} \text{ M}$ $[\text{Ag}^+] = 2.14 \times 10^{-6} \text{ M}$ $[\text{IO}_3^-] = 4.675 \times 10^{-3} \text{ M}$

4. Balance the following unbalanced equation:



5. Consider the following half-reactions:

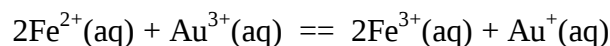


- a) which is the weakest oxidizing agent on the list? **V^{2+}**
 b) Which is the strongest reducing agent? **V(s)**
 c) Will Pb(s) reduce $\text{V}^{2+}(\text{aq})$ to V(s) ? **NO**
 d) Will $\text{I}_2(\text{s})$ oxidize $\text{Cl}^-(\text{aq})$ to $\text{Cl}_2(\text{g})$? **NO**
 e) Name the elements or ions that can be reduced by Pb(s) **$\text{Cl}_2(\text{g})$ and $\text{I}_2(\text{s})$**

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



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



$E^\circ_{\text{cell}} = 0.49 \text{ V}$ $K = 3.8 \times 10^{16}$