Chemistry 350 Spring 2016
20 points Due In Class Wednesday January 20

Problem Set 1 -- Stoichiometry and Concentration Review

Complete all problems on separate paper. Show all work for credit. Correct use of significant figures is required for full credit.

1. Describe how to prepare 2.00 L of a solution that has a chloride concentration of 0.0100 M starting with:
   a. solid copper (II) chloride
   b. 0.200 M copper (II) chloride solution.

a.

$$\frac{0.010\ \text{mol } Cl^-}{1\ \text{L}} \times 2.00\ \text{L} \times \frac{1\ \text{mol } CuCl_2}{2\ \text{mol } Cl^-} \times \frac{134.452\ \text{g } CuCl_2}{1\ \text{mol } CuCl_2} = 1.35\ \text{g } CuCl_2$$

So, weigh 1.35 g $CuCl_2$ and dilute it to 2.00 L.

b.

$$\frac{0.010\ \text{mol } Cl^-}{1\ \text{L}} \times 2.00\ \text{L} \times \frac{1\ \text{mol } CuCl_2}{2\ \text{mol } Cl^-} \times \frac{1\ \text{L}}{0.200\ \text{mol } CuCl_2} = 0.050\ \text{L } CuCl_2$$

So, diluted 0.050 L (50 mL) of 0.200 M $CuCl_2$ ti 2.00 L.

2. When calcium carbonate is added to hydrochloric acid, calcium chloride, carbon dioxide, and water are produced. How many grams of calcium chloride will be produced when 31.0 g of calcium carbonate are combined with 10.0 g of hydrochloric acid?

$$CaCO_3 + 2HCl \rightarrow H_2O + CO_2 + CaCl_2$$

$$31.0\ \text{g } CaCO_3 \times \frac{1\ \text{mol } CaCO_3}{100.067\ \text{g } CaCO_3} \times \frac{1\ \text{mol } CaCl_2}{1\ \text{mol } CaCO_3} \times \frac{110.984\ \text{g } CaCl_2}{1\ \text{mol } CaCl_2} = 34.37\ \text{g } CaCl_2$$

$$10.0\ \text{g HCl} \times \frac{1\ \text{mol HCl}}{36.461\ \text{g HCl}} \times \frac{1\ \text{mol } CaCl_2}{2\ \text{mol HCl}} \times \frac{110.984\ \text{g } CaCl_2}{1\ \text{mol } CaCl_2} = 15.22\ \text{g } CaCl_2$$

So, HCl is the limiting reactant and 15.2 g of $CaCl_2$ will be produced.

3. 0.850 L of 0.490 M sulfuric acid is mixed with 0.800 L of 0.250 M potassium hydroxide. What concentration of sulfuric acid remains after neutralization?

$$H_2SO_4 + 2\ KOH \rightarrow 2H_2O + K_2SO_4$$

$$0.800\ \text{L KOH} \times \frac{0.250\ \text{mol KOH}}{1\ \text{L KOH}} \times \frac{1\ \text{mol } H_2SO_4}{2\ \text{mol KOH}} = 0.100\ \text{mol } H_2SO_4 \text{ consumed}$$

$$0.850\ \text{L } H_2SO_4 \times \frac{0.490\ \text{mol } H_2SO_4}{1\ \text{L } H_2SO_4} = 0.4165\ \text{mol } H_2SO_4 \text{ introduced}$$

0.4165-0.100 = 0.3165 mol H_2SO_4 remain/(0.800+0.850)L = **0.192 M H_2SO_4**

4. In the laboratory, you weigh out 0.114 grams of solid $AlCl_3 \cdot 6H_2O$ (molar mass 241.43 g/mol), dissolve it, and dilute it to a total volume of 100.0 mL to make solution A. You then transfer 3.00 mL of solution A into a 25.0 mL volumetric flask and dilute to the mark to make solution B. What is the concentration of aluminum ion in solution B in moles per liter? In ppm?

Solution A:

$$0.114\ \text{g } AlCl_3 \cdot 6H_2O \times \frac{1\ \text{mol } AlCl_3 \cdot 6H_2O}{241.43\ \text{g } AlCl_3 \cdot 6H_2O} \times \frac{1\ \text{mol } Al^{3+}}{1\ \text{mol } AlCl_3 \cdot 6H_2O} \times \frac{1}{0.1000\ \text{L}} = 4.72 \times 10^{-3}\ \text{M}$$

Solution B:

$$3\ \text{mL} \times \frac{4.72 \times 10^{-3}\ \text{mol } Al^{3+}}{1\ \text{L}} \times \frac{1}{25\ \text{mL}} = 5.67 \times 10^{-4}\ \text{M}$$

In ppm:

$$\frac{5.67 \times 10^{-4}\ \text{mol } Al^{3+}}{1\ \text{L}} \times \frac{26.982\ \text{g Al}}{1\ \text{mol}} \times \frac{10^3\ \text{mg}}{1\ \text{g}} = \frac{15.3\ \text{mg Al}}{1\ \text{L}} = 15.3\ \text{ppm}$$

5. An iron ore sample weighing 0.9132 g is dissolved in HCl(aq), and the iron is obtained as $Fe^{2+}(aq)$. This solution is then titrated with 28.72 mL of 0.05051 M $K_2Cr_2O_7(aq)$ using the balanced reaction below. What is the mass percent of iron in the iron ore?

$$6Fe^{2+} + 14H^+ + Cr_2O_7^{2-} \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$$

$$0.0287\ \text{L} \times \frac{0.05051\ \text{mol } Cr_2O_7^{2-}}{\text{L}} \times \frac{6\ \text{mol Fe}}{1\ \text{mol } Cr_2O_7^{2-}} \times \frac{55.845\ \text{g Fe}}{1\ \text{mol Fe}} = 0.4857\ \text{g Fe}$$

$$\frac{0.4857\ \text{g Fe}}{0.9132\ \text{g ore}} \times 100\% = 53.19\%\ \text{Fe}$$