Post-lab 5

1. 250.0 ml of 0.320 M copper (II) sulfate solution is reacted with 100.0 ml of 1.50 M potassium hydroxide. The *unbalanced* molecular equation for this reaction is:

$$\_\_CuSO_4 + \_\_KOH \rightarrow \_\_Cu(OH)_2 + \_\_K_2SO_4$$

Balance the equation for this reaction and determine the limiting reagent for this reaction using the information given above. Then report the number of moles of $Cu(OH)_2$ that can be formed before the limiting reagent runs out.

1) balance the chemical equation
2) calculate moles of $CuSO_4$ available (using M × V)
3) calculate moles of $Cu(OH)_2$ formed from this (using the balanced equation)

4) calculate moles of KOH (using M × V)
5) calculate moles of $Cu(OH)_2$ formed from this (using the balanced equation)

Whichever one gives you fewer moles of product is your limiting reagent. Report the number of moles formed.

2. Determine the limiting reactant for the following reaction, given that 500 grams of each reactant was used.

$$3\ Cu_{(s)} + 2H_2O_{(l)} + SO_{2(g)} + 2\ O_{2(g)} \rightarrow Cu_3(OH)_4SO_{4(s)}$$

1) For each of the four reactants, calculate the number of moles available using $\text{moles} = \dfrac{500\,\text{grams}}{\text{MW}}$
2) For each of the four reagents, use the balanced equation to determine how many moles of product can be formed with the moles of reactant you got in step 1.
3) Whichever one forms the fewest number of moles of product is your limiting reagent.

3. Antimony sulfide reacts with iron metal according to the following equation:

$$Sb_2S_{3(s)} + 3\ Fe_{(s)} \rightarrow 2\ Sb_{(s)} + 3\ FeS_{(s)}$$

How much pure antimony in kilograms could be produced from 250.0 kg of antimony sulfide and 120.0 kg of iron?

1) Calculate how many moles of each reagent you have using

moles $Sb_2S_3 = \dfrac{\text{grams of } Sb_2S_3}{MW_{Sb_2S_3}}$ and moles Fe $= \dfrac{\text{grams of Fe}}{AW_{Fe}}$.

2) Using the balanced equation, determine how many moles of Sb you can form using the moles of reactants you have available.

3) Whichever one is less is your limiting reagent. Convert that number of moles of Sb to grams of Sb.

4. 1.00 L of pure methanol, CH_3OH, is reacted with 1.00 kg of oxygen according to the following reaction:

$$2\ CH_3OH_{(l)} + 3\ O_{2(g)} \rightarrow 2\ CO_{2(g)} + 4\ H_2O_{(l)}$$

Determine the amount of carbon dioxide in grams that is produced in the reaction, assuming 83.5% yield based on the limiting reactant. The density of methanol is 0.793 g/ml.

First determine the limiting reagent. Then determine the theoretical yield. Then calculate the actual yield.

1) Convert liters of methanol to grams of methanol using the density. Then calculate moles of methanol by dividing by the molecular weight of methanol.

2) Calculate moles of oxygen using the mass and molecular weight of O_2.

3) Using the balanced equation, calculate moles of CO_2 formed from your moles of CH_3OH, and moles of CO_2 formed from your moles of O_2. Whichever one is less is your limiting reagent.

4) Multiply moles of CO_2 formed by the molecular weight of CO_2 to get grams of CO_2. Then multiply by 0.835 to get the actual yield in grams.

5. 250.0 ml of 0.320 M copper (II) sulfate solution is reacted with 100.0 ml of 1.50 M potassium hydroxide and 5.950 grams of solid copper (II) hydroxide is produced. Calculate the percent yield of copper (II) hydroxide.

In Problem 1 we already calculated the number of moles that would be formed based on the limiting reagent.

1) Moles of Cu(OH)2 × $MW_{Cu(OH)2}$ = theoretical yield of $Cu(OH)_2$

2) $\%\ \text{yield} = \dfrac{\text{grams of } Sb_2S_3 \text{ formed (given in the problem)}}{\text{theoretical yield}} \times 100$

More explanations about limiting reagent, theoretical yield, and percent yield can be found in Zumdahl on pages 72- 77.