EXPERIMENT 7 – Reaction Stoichiometry and Percent Yield

INTRODUCTION

Stoichiometry calculations are about calculating the amounts of substances that react and form in a chemical reaction. The word "stoichiometry" comes from the Greek *stoikheion* "element" and *metriā* "measure."

Based on the balanced chemical equation, we can calculate the amount of a product substance that will form if we begin with a specific amount of one or more reactants. Or, you may have a target amount of product to prepare. How much starting compounds are needed to prepare this amount? These are practical calculations that are done frequently by chemists.

In this experiment, you will prepare copper metal from the reaction of aluminum metal with a solution of copper(II) sulfate (cupric sulfate). From the amounts of the reactants, you will determine which reactant is the **limiting reactant**, and from this amount, calculate the **theoretical yield** of copper metal. From the actual amount of copper obtained, you can then calculate your **percent yield** of copper.

THE CALCULATION METHOD

For practically all stoichiometry calculations, we want to use the . . . .

Fabulous Four Steps

Step 1: Write the balanced chemical equation for the reaction.

Step 2: Calculate the *moles* of "given" substance. If more than one reactant amount is given, calculate the moles of each to determine which is the *limiting reactant*.

Step 3: Calculate the moles of "desired" substance from your answer in Step 2 using the coefficients from the balanced chemical equation. If more than one reactant was given originally, you can calculate the moles of product *twice*, based on the moles of each reactant. The reactant that gives the smaller moles of product is the limiting reactant. Keep this answer for Step 4.

Step 4: Convert your answer in Step 3 to the units the problem asks for. Usually this is grams, but it could be volume (for gases or liquid solutions) or concentration (such as molarity, for solutions).

Again, in brief:

1. Balanced reaction.
2. Moles of "given" substance(s).
3. Moles of "desired" substance such as a product.
4. Convert Step 3 answer to the units asked for.

SAMPLE CALCULATIONS

EXAMPLE 1. How many grams of iron(III) oxide (ferric oxide), Fe_2O_3, are formed from the reaction of 5.00 g of iron metal with excess oxygen gas?

Step 1: Balanced reaction.

$$\mathbf{4}\ Fe\ (s) + \mathbf{3}\ O_2\ (g) \longrightarrow \mathbf{2}\ Fe_2O_3\ (s)$$

Step 2: Moles of "given" substance. The given substance is the iron metal, since its amount is given.
moles of Fe = grams of Fe / molar mass of Fe = 5.00 g / 55.845 g/mol = 0.08953353 mol (not rounding yet). Note that gram units cancel, leaving the unit of moles for the answer

$$\text{moles of Fe} = \frac{5.00\ \text{g}\ \cancel{\text{Fe}}}{1} \times \frac{1\ \text{mol Fe}}{55.847\ \text{g}\ \cancel{\text{Fe}}} = 0.08953353\ \text{mol Fe}$$

Step 3: Moles of "desired" substance. This is the product, ferric oxide.
Here is where the balanced reaction comes in. The coefficients in the balanced reaction represent the <u>moles</u> of the substances that react and form. As such, balanced reactions are "in" moles by default. This is why we always have to <u>convert amounts to moles</u> when working stoichiometry problems.

$$\text{moles of } Fe_2O_3 = \frac{0.08953353\ \cancel{\text{mol Fe}}}{1} \times \frac{\mathbf{2}\ \text{mol } Fe_2O_3}{\mathbf{4}\ \cancel{\text{mol Fe}}} = 0.044766765\ \text{mol } Fe_2O_3$$

Step 4: Grams of Fe_2O_3.

Grams of Fe_2O_3 = moles of Fe_2O_3 X molar mass of Fe_2O_3

$$= 0.044766765\ \cancel{\text{mol}} \times 159.6882\ \text{g}/\cancel{\text{mol}} = \boxed{\mathbf{7.15\ g}}$$ rounded to 3 significant figures.

And that's it! We can also do steps 2 through 4 like a conversion problem. Once you are familiar with the steps, you can work these problems more quickly this way:

$$\frac{5.00\ \text{g}\ \cancel{\text{Fe}}}{1} \times \frac{1\ \cancel{\text{mol Fe}}}{55.845\ \cancel{\text{g Fe}}} \times \frac{2\ \text{mol}\ \cancel{Fe_2O_3}}{4\ \cancel{\text{mol Fe}}} \times \frac{159.6882\ \text{g } Fe_2O_3}{\cancel{\text{mol } Fe_2O_3}} = \mathbf{7.15\ g}\ \text{of } Fe_2O_3$$

When working stoichiometry problems, always do the following:

a) Write the <u>units</u> on all numbers.
b) Check that the <u>units cancel properly</u>.
c) Give the <u>correct unit for the answer</u>.
d) <u>Avoid rounding numbers too much</u> during the calculation, or you will have <u>roundoff error</u> in your answer.
e) Round your <u>final answer</u> to the <u>correct number of significant figures</u>.

EXAMPLE 2. If 10.0 g of iron metal is reacted with 15.0 g of Cl_2 gas, how many grams of ferric chloride, $FeCl_3$, will form?

In this problem, the amounts of both reactants are given, so we will have to determine which reactant is the limiting reactant (the one that "limits" the amount of product that is formed). The other reactant is in excess amount. We'll use the Fab Four Steps just as before.

Step 1: Balanced reaction.

$$2\ Fe\ (s)\ +\ 3\ Cl_2\ (g) \longrightarrow 2\ FeCl_3\ (s)$$

Step 2: Moles of given substances.

moles of Fe = 10.0 g / 55.845 g/mol = 0.17906706 mol

moles of Cl_2 = 15.0 g / 70.906 g/mol = 0.211547682 mol

Step 3: Moles of desired substance, $FeCl_3$.

Since we have two given amounts, a straightforward approach to this step is to calculate the moles of $FeCl_3$ twice, first based on the moles of Fe and second based on the moles of Cl_2. Keep the *smaller* answer. The reactant that gives this smaller answer is the *limiting reactant*. The other reactant is in excess amount.

$$\text{moles of } FeCl_3 \text{ based on Fe} = \frac{0.17906706 \text{ mol Fe}}{1} \times \frac{2 \text{ mol } FeCl_3}{2 \text{ mol Fe}} = 0.17906706 \text{ mol of } FeCl_3$$

$$\text{moles of } FeCl_3 \text{ based on } Cl_2 = \frac{0.211547682 \text{ mol } Cl_2}{1} \times \frac{2 \text{ mol } FeCl_3}{3 \text{ mol } Cl_2} = 0.141031788 \text{ mol of } FeCl_3$$

Keep this answer!

Since the moles of $FeCl_3$ based on moles of Cl_2 is the smaller answer, Cl_2 is the limiting reactant. Iron metal is therefore in excess amount, so there will be some Fe left over as unreacted excess.

Note that we might have reasonably assumed that iron metal was the limiting reactant since it was present in lesser amount in grams initially (10.0 g of Fe and 15.0 g of Cl_2). But it turned out that Cl_2 was the limiting reactant. The molar masses of the substances and the reaction stoichiometry come into play also, so we can't automatically assume which substance is the limiting reactant until we go through the steps as we did above.

Step 4: Finally! Convert moles of $FeCl_3$ to grams.

Grams of $FeCl_3$ = 0.141031788 mol X 162.204 g/mol = $\boxed{\textbf{22.9 g}}$ rounded to 3 significant figures.

Practice makes perfect! Working stoichiometry problems properly will strengthen your skills in working many other types of chemistry problems as well.

PERCENT YIELD

The percent yield of a reaction tells us how well the reaction worked in terms of forming a desired product.

$$\textbf{Percent Yield} = \frac{\textbf{Actual (Experimental) Yield of Product}}{\textbf{Theoretical (Calculated) Yield of Product}} \times 100$$

Remember, "Actual Over Theoretical Times 100."

The unit of the amounts may be in grams or moles. Usually we use grams, since we will have weighed the product on a balance in gram units.

EXAMPLE. In Example 1 above, let's say that the student obtained 6.75 g of ferric oxide product. The theoretical yield of ferric oxide calculated with the Fabulous Four Steps was 7.15 g. The percent yield of ferric oxide from the reaction is therefore,

$$\textbf{Percent Yield of } \mathbf{Fe_2O_3} = \frac{\textbf{6.75 g}}{\textbf{7.15 g}} \times \textbf{100} = \boxed{\textbf{94.4\%}}$$

Generally, less than 100% yields are obtained. The reaction may not have sufficient time to go to completion, and if the product is transferred from on container to another, or is filtered, there is always some small loss. Therefore, the less "handling," the better the yield will be. If undesired "side reactions" occur in addition to the desired reaction, the yield will be affected. The use of impure reactants can adversely affect the percent yield also.

On the other hand, it is impossible to obtain <u>more</u> than a 100% yield! If your calculated percent yield is more than 100%, then, aside from weighing errors, there is some additional substance present in your product making it weigh more than it should. If a solid product is prepared in aqueous solution, <u>incomplete drying</u> is the usual culprit, since the wet solid, which may <u>appear</u> dry, will weigh more than the dry solid product. Chemists use the method of <u>heating to constant weight</u> to avoid this error.

If other impurities are present in the product, other methods must be used to remove them, such as <u>extraction</u>, <u>recrystallization</u>, <u>distillation</u>, and <u>chromatography</u>. These techniques are often used by chemists to isolate and purify the desired product from the reaction mixture as effectively and efficiently as possible.

EXPERIMENTAL PROCEDURE

★ **<u>Safety Note: Hydrochloric acid is a strong acid that is harmful to the skin and especially to your eyes. Wear your safety glasses or goggles during the entire procedure. The reaction also produces flammable hydrogen gas (H_2), so Bunsen burners should not be used while the reaction is in progress.</u>**

Also see the procedural notes on the next page.

1) Weigh a clean, dry 150 mL beaker and record its weight on the report form.

2) Carefully add copper(II) sulfate pentahydrate, $CuSO_4 \cdot 5\ H_2O$, until 2.00 g have been added.

3) Measure 10 mL of deionized water in a small graduated cylinder and add the water to the beaker to dissolve the copper(II) sulfate pentahydrate with the aid of a glass stirring rod. Record the color of the solution on the report sheet.

4) Measure 2.0 mL of 6 M HCl in your graduated cylinder, add it to the solution, and mix well.

5) Weigh 0.25 g of dry aluminum foil in small pieces and record the weight on the report sheet.

6) Add the pieces of Al foil a little at a time. Use the stirring rod to mix the solution during the reaction. (CAUTION: Exothermic reaction!) Note the color of the solution after the added piece of aluminum no longer darkens on its surface. Add the remaining few pieces of aluminum foil, and add an additional 5 mL of 6 M HCl to facilitate the reaction of any excess aluminum with the hydrochloric acid.

7) After all of the aluminum foil has reacted, allow the solid particles of copper product to settle, and carefully decant the solution from the solid (leaving the copper behind in the beaker). Add 20 mL of deionized water to the solid, stir well with the stirring rod, and decant again. Repeat this washing with 20 mL of water once more. Finally, add 10 mL of methanol to the solid, stir, and decant.

8) Heat the beaker on a electric hot plate at medium heat (a setting of about 4 out of 10) until the solid and beaker are thoroughly dry. Allow the beaker and its contents to cool, and then weigh and record the weight on the report form. If time allows, you may re-heat the beaker for an additional 10 minutes and re-weigh the beaker and contents to ensure that drying is complete (heating to constant weight).

9) On the report sheet, record the moles of Al and $CuSO_4 \cdot 5\ H_2O$ used, and determine which is the limiting reactant. Based on the limiting reactant, calculate the theoretical yield of copper metal product. From your actual yield of copper, calculate the percent yield of copper product obtained from the reaction.

10) Place your copper metal in the collection container on the front desk (do not wash it down the sink!), rinse your glassware well, and return it and your other equipment to their proper storage locations.

PROCEDURAL NOTES

Step 2. A hydrate is a solid compound that contains "trapped" water molecules in the solid. In copper(II) sulfate pentahydrate, one mole of the solid $CuSO_4$ has 5 moles of water molecules trapped in it. The water molecules are included in the molar mass of $CuSO_4 \cdot 5\,H_2O$, which is **249.69 g/mol**.

Step 4. Chloride ion facilitates the reaction of Al with $CuSO_4 \cdot 5\,H_2O$. A convenient source of chloride ion is from hydrochloric acid, and in addition, the hydrochloric acid will react with any excess aluminum, allowing it to be separated from the solid copper product. The relevant reactions are,

a) The reaction of aluminum with copper(II) sulfate in aqueous solution. This reaction ordinarily occurs rather slowly, but when chloride ion is present, the reaction rate increases dramatically.

$$2\,Al\,(s) \;+\; 3\,CuSO_4\,(aq) \longrightarrow 3\,Cu\,(s) \;+\; Al_2(SO_4)_3\,(aq)$$

b) The competing reaction of aluminum with hydrochloric acid. This reaction still occurs more slowly than the above chloride-aided reaction with the reactant concentrations that are used. This reaction thus has a minimal effect on the percent yield, so the maximum amount of copper should be able to form if excess aluminum is used.

$$2\,Al\,(s) \;+\; 6\,HCl\,(aq) \longrightarrow 3\,H_2\,(g) \;+\; 2\,AlCl_3\,(aq)$$

Both of these reactions are <u>single replacement</u> reactions. Al is a more active element than Cu in the first reaction, and in the second reaction, Al is more active than the element H_2 (recall the <u>activity series</u>:
Li > K > Ba > Sr > Ca > Na > Mg > Al > Mn > Zn > Fe > Cd > Co > Ni > Sn > Pb > H_2 > Cu > Ag > Hg > Au).
On the other hand, as seen from the series, the element Cu is <u>less</u> active than H_2, so the copper metal product will <u>not</u> react with the hydrochloric acid that is present:

$$Cu\,(s) \;+\; 2\,HCl\,(aq) \;\not\longrightarrow\; H_2\,(g) \;+\; CuCl_2\,(aq)$$

Step 7. The methanol (methyl alcohol) wash removes additional water, facilitating complete drying of the product.

Step 8. Heating the copper particles at too high of a temperature in air results in an undesirable darkening and a gain in weight, apparently due to the formation of black cupric oxide, CuO. During heating, the dry particles of copper should remain loose when stirred and should not darken in color.

EXPERIMENT 7 – Reaction Stoichiometry and Percent Yield

REPORT FORM

Name ______________________________

Instructor ___________________________

Date _______________________________

1. Mass of empty 150 mL beaker ______________ g

2. Mass of beaker plus $CuSO_4 \cdot 5\ H_2O$ ______________ g

3. Color of solution ______________

4. Mass of $CuSO_4 \cdot 5\ H_2O$ [2] – [1] ______________ g

5. Mass of Al foil used (small pieces) ______________ g

6. Color of solution after reaction is complete ______________

7. Mass of beaker and copper product (after heating and cooling)

First weighing ______________ g

Second weighing ______________ g

8. Mass of copper metal product [7] – [1] ______________ g

9. Moles of Al used ______________ mol
(show your calculation)

10. Moles of $CuSO_4 \cdot 5\ H_2O$ used ______________ mol
(show your calculation)

11. Moles of copper product based on moles of Al ______________ mol
(show your calculation)

12. Moles of copper product based on moles of $CuSO_4 \cdot 5\ H_2O$ ______________ mol
(show your calculation)

13. Limiting reactant ______________

14. Grams of Cu product based on the limiting reactant (theoretical yield) ______________ g
(show your calculation)

15. Percent yield of Cu (2 to 3 significant figures) ______________ %
(show your calculation)

EXPERIMENT 7 **Name**

Pre-Laboratory Questions and Exercises

Due before lab begins. Answer in the space provided.

1. What safety precautions are cited in this experiment?

2. What it the purpose of adding hydrochloric acid to the reaction mixture?

3. A student combusted 0.500 g of purified aluminum powder with excess oxygen in an oxygen atmosphere according to the reaction,

$$4\ Al\ (s) + 3\ O_2\ (g) \longrightarrow 2\ Al_2O_3\ (s)$$

a) What is the limiting reactant? ____________

b) How many moles of Al were used? ____________ mol of Al

c) How many moles of Al_2O_3 would form, based on the moles of the limiting reactant? ____________ mol of Al_2O_3

d) How many moles of O_2 are required to react completely with the aluminum? ____________ mol of O_2

e) How many grams of O_2 are required to react completely with the aluminum? ____________ g of O_2

f) What is the theoretical yield of Al_2O_3 in grams? ____________ g of Al_2O_3

g) If the student collected 0.918 grams of Al_2O_3 product, what was the percent yield of Al_2O_3 obtained? ____________ %

EXPERIMENT 7 **Name**

Post-Laboratory Questions and Exercises

Due after completing the lab. Answer in the space provided.

1. Heating the copper product at too high a temperature in an oxygen atmosphere results in the formation of copper (II) oxide, or cupric oxide, CuO. Write the balanced chemical equation for this reaction.

2. What are some reasons for obtaining a percent yield of less than 100 percent?

3. A student reacted 0.500 g of $Pb(NO_3)_2$ with 0.750 g of KI according to the reaction,

$$Pb(NO_3)_2\ (aq) + 2\ KI\ (aq) \longrightarrow PbI_2\ (s) + 2\ KNO_3\ (aq)$$

a) How many moles of $Pb(NO_3)_2$ were used? ____________ mol of $Pb(NO_3)_2$

b) How many moles of KI were used? ____________ mol of KI

c) How many moles of PbI_2 would form, based on the moles of $Pb(NO_3)_2$ used? ____________ mol of PbI_2

d) How many moles of PbI_2 would form, based on the moles of KI used? ____________ mol of PbI_2

e) Which is the limiting reactant? ____________

f) What is the theoretical yield of PbI_2 in grams? ____________ g of PbI_2

g) If the student obtained 0.583 grams of PbI_2 product after collecting it by filtration and drying it, what was the percent yield of PbI_2 obtained? ____________ %