Experiment 12
Stoichiometry

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Name:

Date:

Key Objectives

1. Calculate Molecular Weight (MW).
2. Calculate Limiting Reactant (LR).
3. Calculate Excess Reactant left over (ER).
4. Calculate Theoretical Yield.
5. Calculate Percent Yield.

Discussion

A common chemical question is how much of a product can I make, or how much of a reactant do I need in a chemical reaction. In this lab you will learn how to answer both of these questions and a few more!

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.

Kitchen Chemistry

Chemistry and eating have a lot in common (ok its a stretch). In Figure **??** you have 10 wieners + 8 buns and can make 8 hot dogs but you will have 2 left over weiners.

Figure 12.1: Why do wieners come in packages of 10 but the buns in a package of 8. That leaves 2 left over wieners. credit: author

In the previous example it was easy to do the math, lets take a look at a slightly harder example and include some math. Imagine you are making a sandwich using the following recipe:

$$\underline{\ 2\ }\,\text{slice bread} + \underline{\ 3\ }\,\text{slices ham} + \underline{\ 3\ }\,\text{slices turkey} \longrightarrow \underline{\ 1\ }\,\text{great sandwich} \tag{25}$$

If you have 12 slices of bread, 24 slices of ham and 36 slices of turkey how many sandwich's can you make? Since this is real life it seems easier than chemistry (hint: it might be, but you can take your experience in real life and apply it to chemistry!). What you (probably without thinking of it) would determine is how many sandwich's you can make with each ingredient: 12 slices of bread you can make 6 sandwich's, with 24 slices of ham you can make 8 sandwich's and with 36 slices of turkey you could make 12 sandwich's. Thus you can make 6 sandwich's before you run out of bread. That makes bread your **Limiting Reactant (LR)**! In other words, the reactant you run out of in a reaction is the Limiting Reactant.

If we imagined this as a chemist would you might do something like the following and deduce that the smallest number of sandwich's made tells us the Limiting Reactant.

$$\frac{12\ \text{slices bread}}{} \times \frac{1\ \text{sandwich}}{2\ \text{slices bread}} = 6\ \text{sandwich}'\text{s} \tag{26}$$

$$\frac{24\ \text{slices ham}}{} \times \frac{1\ \text{sandwich}}{3\ \text{slices ham}} = 8\ \text{sandwich}'\text{s} \tag{27}$$

$$\frac{36\ \text{slices turkey}}{} \times \frac{1\ \text{sandwich}}{3\ \text{slices turkey}} = 12\ \text{sandwich}'\text{s} \tag{28}$$

We can now define a few more terms:

Excess Reactant (ER): The reactant that you have excess (or left over of).

Theoretical Yield: The maximum amount of a product you can make based on calculations.

Actual Yield: The actual yield of a product in the laboratory (or kitchen), What you actually measure/make.

Percent Yield: Is the percentage of the theoretical value you make in the lab or kitchen. It is essentially your "grade".

Mathematically we can write:

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100 \tag{29}$$

If we take our sandwich making example then we could say that the we have **Excess** ham (6 slices) and turkey (18 slices) and we can **Theoretically** make 6 sandwich's. If we dropped a slice of bread (yuck, no 5 second rule in chemistry lab) then our **Yield** of sandwich's would go down because we can only make 5 sandwich's.

Again switching to thinking like a chemist we might formulate the problem as follows:

To find the Excess Reactant left over we start with Limiting Reactant and calculate how much of the Excess Reactant we use and subtract it from what we started with to get the amount left over.

$$\frac{12\ \text{slices bread}}{} \times \frac{3\ \text{slices ham}}{2\ \text{slices bread}} = 18\ \text{slice}'\text{s ham} \therefore 24 - 18 = 6\ \text{slice}'\text{s ham left over} \tag{30}$$

$$\frac{12\ \text{slices bread}}{} \times \frac{3\ \text{slices turkey}}{2\ \text{slices bread}} = 18\ \text{slice}'\text{s turkey} \therefore 36 - 18 = 18\ \text{slice}'\text{s turkey left over} \tag{31}$$

We have already calculated our Theoretical Yield of sandwich's when we calculated the LR in equation 2-4 and found we can only make 6 sandwich's

Our Percent Yield (assuming we dropped a slice of bread would be:

$$\text{Percent Yield} = \frac{5\ \text{sandwich}'\text{s}}{6\ \text{sandwich}'\text{s}} \times 100 = 83\% \tag{32}$$

Mol to Mol ratio

When making sandwich's we know the amount of ingredients to use, so to in chemistry we have a chemical reaction that tells us the amount of each chemical to use. Instead of slices chemists use a mole (or mol for short) which is just a ridiculously large number of slices. 6.02×10^{23} to be exact. To give credit where it is due, Avagadro discovered this and so its generally referred to as Avagadro's number: 1 mol = 6.02×10^{23} anything.

For our sandwich recipe we might write the reactions below. The first reaction we are measuring everything by the "slice". In the second reaction we use the chemistry term "mol" which just means lots and lots and lots of slices. In the third recipe/reaction we omit the word "mol" because we are chemists and we know we measure everything in "mols"

2	slice bread	+	3	slice ham	+	3	slice turkey	$\longrightarrow$	1	sandwich
2	mol bread	+	3	mol ham	+	3	mol turkey	$\longrightarrow$	1	mol sandwich
2	bread	+	3	ham	+	3	turkey	$\longrightarrow$	1	sandwich

Thus, we can have a slice/slice ratio if making sandwich's and a mol/mol ratio if we are in the chemistry lab.

$$\frac{1\ \text{sandwich}}{3\ \text{slices ham}} \text{ or } \frac{1\ \text{mol sandwich}}{3\ \text{mol ham}} \tag{33}$$

Molecular Weight (MW)

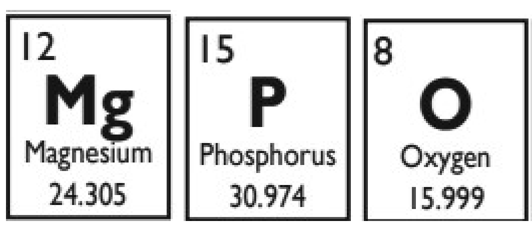
If we went to the Deli to buy sandwich meat to make our sandwich's we would probably buy everything by the ounce (oz). If we did this then we might need to know how many slices of ham are in each ounce of ham we order (same goes for turkey). This would then be a conversion factor, for example we might say that 1 ounce of ham = 6 slices, but turkey is sliced thicker and 1 ounce of turkey = 3 slices. Thus, to order the same number of ham and turkey slices we would need to buy twice as many ounces of turkey as of ham. The number of ham/turkey slices and the mass of the ham/turkey are different.

$$\frac{1\text{ oz ham}}{6\text{ slices ham}}\text{ and }\frac{1\text{ oz turkey}}{3\text{ slices turkey}} \tag{34}$$

If we went to the store and wanted to buy 24 slices of ham we could calculate how many ounces we needed:

$$\frac{24\text{ slices ham}}{} \times \frac{1\text{ oz ham}}{6\text{ slices ham}} = 4\text{ oz ham} \tag{35}$$

In the kitchen we might use cups or ounces to measure things but in the chemistry lab we measure out our chemicals in grams. Just like our ham and turkey each chemical compound has a different mass. In order to get from the grams that we measure in lab to the moles that our balanced equation is written in we will need to calculate the Molecular Weight (MW) or it is sometimes referred to as the Molar Mass. This is the weight of 1 mole (6.02×10^{23} atoms or molecules) of a substance and is calculated using the weight of 1 mole of each element on the periodic table.



For example we can calculate the MW of $Mg_3(PO_4)_2$:

$$\begin{aligned}
\text{Mg} &= 3 \times 24.305 &&= 72.915\text{ g/mol}\\
\text{P} &= 2 \times 30.974 &&= 61.948\text{ g/mol}\\
\text{O} &= 8 \times 15.999 &&= 127.992\text{ g/mol}\\
\hline
&\text{Mg}_3(\text{PO}_4)_2 &&= 262.853\text{ g/mol}
\end{aligned} \tag{36}$$

General Procedure

Generically we can apply the following procedure to calculate our Limiting Reactant (LR), Excess Reactant left over (ER), Theoretical and Percent Yields.

Quantity	Calculation
Limiting Reactant	$\mathrm{g\,R_1 \longrightarrow mol\,R_1 \longrightarrow mol\,P_1 \longrightarrow g\,P_1}$ $\mathrm{g\,R_2 \longrightarrow mol\,R_2 \longrightarrow mol\,P_1 \longrightarrow g\,P_1}$ The Limiting Reactant (LR) is the reactant that produces the least amount of Product.
Excess Reactant	$\mathrm{g\,LR \longrightarrow mol\,LR \longrightarrow mol\,ER \longrightarrow g\,ER}$ The Excess Reactant is found by subtraction: Start g ER - Used g ER = Left Over ER
Theoretical Yield	Already calculated! The minimum amount of product made when you calculated your LR.
Percent Yield	$\frac{\text{Actual}}{\text{Theoretical}} \times 100$

Chemistry Example

We can now graduate from the kitchen to the laboratory with one good example below. Given the reaction below assuming you start with 25.0 g of H_2SO_4 and 50.0 g H_3PO_4.

$$\begin{array}{ccccccc} \underline{\ 3\ }\,\mathrm{Mg(OH)_2(aq)} & + & \underline{\ 2\ }\,\mathrm{H_3PO_4} & \longrightarrow & \underline{\ 1\ }\,\mathrm{Mg_3(PO_4)_2(s)} & + & \underline{\ 6\ }\,\mathrm{H_2O(l)} \\ 58.32\ \mathrm{g/mol} & & 98.00\ \mathrm{g/mol} & & 262.85\ \mathrm{g/mol} & & 18.02\ \mathrm{g/mol} \end{array} \tag{37}$$

Below each reactant and product is the molecular weight which we will need in our calculations. You can double check my calculations to make sure that you remember how to do it!

We will answer the following 5 questions (given that you start with 25.0 g of H_2SO_4 and 50.0 g H_3PO_4:

1. Calculate the Limiting Reactant (LR).
2. Calculate the Theoretical Yield of $Mg_3(PO_4)_2$.
3. Calculate the Theoretical Yield of H_2O.
4. Calculate the amount of Excess Reactant left over.
5. Calculate the Percent Yield if in lab you produced 25.0 g of $Mg_3(PO_4)_2$.

Calculating the Limiting Reactant by determining how much of a product can be produced with each reactant. The reactant that produces the least product is your limiting reactant, and as a bonus, you also get the theoretical yield of that product (a 2 for 1 deal!).

$$\frac{25.0\,\mathrm{g\,Mg(OH)_2}}{} \times \frac{1\,\mathrm{mol\,Mg(OH)_2}}{58.32\,\mathrm{g\,Mg(OH)_2}} \times \frac{1\,\mathrm{mol\,Mg_3(PO_4)_2}}{3\,\mathrm{mol\,Mg(OH)_2}} \times \frac{262.85\,\mathrm{g\,Mg_3(PO_4)_2}}{1\,\mathrm{mol\,Mg_3(PO_4)_2}} = 37.6\,\mathrm{g\,Mg_3(PO_4)_2}$$

$$\frac{50.0\,\mathrm{g\,H_3PO_4}}{} \times \frac{1\,\mathrm{mol\,H_3PO_4}}{98.00\,\mathrm{g\,H_3PO_4}} \times \frac{1\,\mathrm{mol\,Mg_3(PO_4)_2}}{2\,\mathrm{mol\,H_3PO_4}} \times \frac{262.85\,\mathrm{g\,Mg_3(PO_4)_2}}{1\,\mathrm{mol\,Mg_3(PO_4)_2}} = 67.1\,\mathrm{g\,Mg_3(PO_4)_2}$$

From the calculations above the least amount of product (37.56 g $Mg_3(PO_4)_2$ which is our Theoretical

Yield and is produced by $Mg(OH)_2$ which is the Limiting Reactant. We have answered question 1 and 2.

To determine theoretical yield of our other product we will start with our Limiting Reactant and determine the amount of water produced.

$$\frac{25.0\,\mathrm{g\,Mg(OH)_2}}{} \times \frac{1\,\mathrm{mol\,Mg(OH)_2}}{58.32\,\mathrm{g\,Mg(OH)_2}} \times \frac{6\,\mathrm{mol\,H_2O}}{3\,\mathrm{mol\,Mg(OH)_2}} \times \frac{18.02\,\mathrm{g\,H_2O}}{1\,\mathrm{mol\,H_2O}} = 15.5\,\mathrm{g\,H_2O}$$

To calculate the amount of Excess Reactant left over we start with our Limiting Reactant, calculate the amount of the other reactant used and subtract it from our starting amount.

$$\frac{25.0\,\mathrm{g\,Mg(OH)_2}}{} \times \frac{1\,\mathrm{mol\,Mg(OH)_2}}{58.32\,\mathrm{g\,Mg(OH)_2}} \times \frac{2\,\mathrm{mol\,H_3PO_4}}{3\,\mathrm{mol\,Mg(OH)_2}} \times \frac{98.00\,\mathrm{g\,H_3PO_4}}{1\,\mathrm{mol\,H_3PO_4}} = 28.0\,\mathrm{g\,H_3PO_4}$$

Start - Used = Left Over
50.0 - 28.0 = 22.0 g H_3PO_4 Left Over

The last calculation is the Percent Yield.

$$\%\,\mathrm{Yield} = \frac{\mathrm{Actual}}{\mathrm{Theoretical}} \times 100$$
$$\%\,\mathrm{Yield} = \frac{25.0\,\mathrm{g\,Mg_3(PO_4)}}{37.6\,\mathrm{g\,Mg_3(PO_4)_2}} \times 100 = 66.5\%$$

To summarize our answers for the 5 questions we posed:

1. Calculate the Limiting Reactant (LR). **$Mg(OH)_2$**
2. Calculate the Theoretical Yield of $Mg_3(PO_4)_2$. **37.6 g $Mg_3(PO_4)_2$**
3. Calculate the Theoretical Yield of H_2O. **15.5 g H_2O**
4. Calculate the amount of Excess Reactant left over. **28.0 g H_3PO_4**
5. Calculate the Percent Yield if in lab you produced 25.0 g of $Mg_3(PO_4)_2$. **66.5 %**

Procedure

In this experiment, you will prepare copper metal (Cu) from the reaction of aluminum metal (Al) with a solution of copper(II) sulfate ($CuSO_4$). 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 reaction we will be studying is given below:

$$\underline{\ 2\ }\ \mathrm{Al(s)} + \underline{\ 3\ }\ \mathrm{CuSO_4(aq)} \longrightarrow \underline{\ 3\ }\ \mathrm{Cu(s)} + \underline{\ 1\ }\ \mathrm{Al_2(SO_4)_3(aq)}$$

1. Weigh a clean, dry 150 mL beaker and record its weight.
2. Weigh out approximately 2.0 g of $CuSO_4$ into the beaker.
3. Add approximately 10. mL of water to the beaker and swirl to dissolve the $CuSO_4$. Record the color of the solution.
4. Measure approximately 2.0 mL of 6 M HCl and add it to the beaker
5. Weigh out approximately 0.25 g of aluminum foil in small pieces and record the mass.
6. Add the pieces of Al foil to the beaker containing the $CuSO_4$ little at a time, stirring the mixture continuously. Use caution as the reaction is very Exothermic. After all the aluminum foil is added to the beaker add an additional 5.0 mL of 6 M HCl to the beaker to facilitate the reaction of any excess Al foil. Record the color of the solution and the solid precipitate.
7. After the reaction is complete allow the solid particles of copper to settle and carefully decant the solution from the solid (leaving the copper behind in the beaker) into a 2nd beaker.
8. Add 20 mL of water to the beaker containing the copper stir well and decant again. Do this step twice.
9. Add 10 mL of methanol to the copper, stir and decant.
10. Dispose of liquids decanted into the waste container labeled "Experiment 14 Waste".
11. Heat the beaker on a hot plate at a medium heat setting (4 out of 10) until the solid and beaker are thoroughly dry. Allow the beaker to cool (so you can touch it) and weight it on the scale. Record the mass.
12. If time permits, reheat the beaker for an additional 10 minutes and record the mass to ensure that the sample was completely dry. This is called heating to a constant mass.
13. Dispose of your copper metal in the container labeled "Solid Copper Metal Waste"

Dispose of any excess materials in the Heavy Metals waste jug (if liquid) or in the provided container (if solid).

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Results

1. Mass of 150 mL beaker 1. ______________
2. Mass of 150 mL beaker + $CuSO_4$: 2. ______________
3. Mass of $CuSO_4$ used: 3. ______________
4. Color of solution (before adding Al foil) 4. ______________
5. Mass of Al foil: 5. ______________
6. Color of solution (after adding Al foil) 6. ______________
7. Mass of dry Cu + beaker (after 1st heating): 7. ______________
8. Mass of dry Cu + beaker (after 2nd heating): 8. ______________
9. Mass of dry Cu: 9. ______________

Post Lab Questions

1. Write the complete (balanced, and include states) chemical reaction that occurred in the lab.

2. Below each Reactant and Product in the previous reaction write the molecular weight of the compound.

3. Determine the theoretical amount of Copper metal from the starting values for the mass of $CuSO_4$. (Show calculation) 3. ______________

4. Determine the theoretical amount of Copper metal from the starting values for the mass of Al. (Show calculation) 4. ______________

5. What is the Theoretical Yield of Copper metal. Explain. 5. ______________

6. What is the Limiting Reactant. Explain. 6. ______________

7. What is the Percent Yield of Copper? (Show calculation) 7. ______________

8. What is the amount of Excess Reactant Left over? (Show calculation) 8. ______________

9. What is the theoretical mass of the $Al_2(SO_4)_3$ produced? (Show calculation.) 9. ______________

10. Did you obey Lavoisier Conservation of Mass law? Explain. 10. ______________

Name: ______________________________ Class: _______________ Date: _______________

Prelab Questions

1. Complete the following table.

Name	**Formula**	**Molecular Weight**
Aluminum Sulfate		
Copper (II) Sulfate	$CuSO_4$	
	HCl	
	$AlCl_3$	
Methanol	CH_3OH	

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