Chapter 9, Section 1

Chemical reactions are what chemistry is all about!

How do you understand a reaction--- Study the equation for a reaction!

What information is given to you in a chemical equation?

- Reactants- substances present before the reaction takes place and the state of matter they are in
- Products- substances present after the reaction takes place and the state of matter they are in

Example of an unbalanced equation: $CO (g) + H_2 (g) \rightarrow CH_3OH (l)$

Tells us that carbon dioxide gas reacts with hydrogen gas to form methanol (CH_3OH)

Since chemical equations represent chemical reactions and chemical reactions are just rearranging atoms, not creating or destroying them, BALANCING chemical equations is necessary. So, What information is given to you in a BALANCED chemical equation?

- Reactants- substances present before the reaction takes place
- Relative amount of each reactant
- Products- substances present after the reaction takes place
- Relative amount of each product

Example of a balanced equation: $CO (g) + 2H_2 (g) \rightarrow CH_3OH (l)$

Tells us that 1 unit of carbon dioxide gas reacts with 2 units of hydrogen gas to form 1 unit of methanol (CH_3OH).

The subscripts tell us relative amounts, these relative amounts can be atoms/molecules (depending on if we are talking about elements or compounds), or moles which represents a very large, but standard (6.022×10^{23}) amount of atoms or molecules.

Because the coefficients represent RELATIVE numbers we are allowed to multiply the ENTIRE equation by any number and we will still have a balanced equation.

Example:

$12 [CO (g) + 2H_2 (g) \rightarrow CH_3OH (l)] = 12 CO (g) + 22 H_2 (g) \rightarrow 12 CH_3OH (l)$

$1/2 [CO (g) + 2H_2 (g) \rightarrow CH_3OH (l)] = 1/2 CO (g) + 1 H_2 (g) \rightarrow 1/2 CH_3OH (l)$

Chapter 9, Section 2

A balanced chemical equation allows us to determine relationships between moles of reactants and moles of products.

Uses of Mole Relationships

- Predict how much of a product will be created if a certain amount of reactant is used.
- Predict how much of a reactant will be necessary to produce a certain amount of product.

Consider the decomposition of water

$2H_2O\ (l) \rightarrow 2H_2\ (g) + O_2\ (g)$

This balanced chemical reaction tells us that for every 2 moles of H_2O that decomposes, 2 moles of H_2 and 1 mole of O_2 will be produced.

Using this reaction, let's examine the two uses of mole relationships:

Predicting how much product will be created from a given amount of reactant

If I have 8 moles of water, how much H_2 and O_2 can I create?

To answer this we have to multiply the entire equation by 4 so that we have the prescribed 8 moles of water.

$[2H_2O\ (l) \rightarrow 2H_2\ (g) + O_2\ (g)] \times 8 = 8H_2O\ (l) \rightarrow 8H_2\ (g) + 4O_2\ (g)$

We can now answer the question: I can create 8 moles of H_2 and 4 moles of O_2

What if I needed to create 20 moles of O_2 for an experiment using the decomposition of water as my source of oxygen. How much water would I need?

Again, I need to multiply my original equation by 20 to obtain the necessary amount of oxygen.

$[2H_2O\ (l) \rightarrow 2H_2\ (g) + O_2\ (g)] \times 20 = 40H_2O\ (l) \rightarrow 40H_2\ (g) + 20O_2\ (g)$

Now I can answer my question, I would need 40 moles of water to create 20 moles of oxygen.

This method works for simple problems such as the ones we just did, however in life there are not many simple whole number problems. For more complex problems we need to use **MOLE RATIOS.**

MOLE RATIOS: are conversion factors, based on the balanced chemical equation, that are used to convert relative amounts of reactant to product or vice versa.

Using the decomposition of water again, we can determine the mole ratios.

$2H_2O\ (l) \rightarrow 2H_2\ (g) + O_2\ (g)$

From the above balanced equation the following equivalency statements can be made:

1) 2 moles of H_2O = 2 moles of H_2
2) 2 moles of H_2O = 1 mole of O_2

These are not literal equalities, just chemical equivalencies.

The equivalency statements can then be used like equality statements to develop conversion factors.

If I have 5.8 moles of water how much hydrogen gas and oxygen gas will I create by decomposition of all the water?

5.8 moles of water	2 moles of hydrogen
	2 moles of water

5.8 ~~moles of water~~	2 moles of hydrogen	= 5.8 * 2/2 = 5.8 * 1 = 5.8 moles of Hydrogen
	2 ~~moles of water~~	

Solving the above problem for amount of hydrogen produced from the decomposition of 5.8 moles of water, yields the answer of 5.8 moles of hydrogen.

Next we solve for the amount of oxygen produced in the same decomposition reaction.

5.8 moles of water	1 mole of oxygen
	2 moles of water

5.8 ~~moles of water~~	1 moles of oxygen	= 5.8 * 1/2 = 5.8 * 1/2 = 2.9 moles of Oxygen
	2 ~~moles of water~~	

This yields 2.9 moles of oxygen.

So the final answer is that if 5.8 moles of water are decomposed, 5.8 moles of Hydrogen gas and 2.9 moles of oxygen gas will be produced from the reaction.

This is where cancellation of units becomes vitally important!!

Remember when putting the work into your calculator, multiply all numbers on the top row of your T- chart, then divide by all numbers on the bottom of your chart!

Try a few:

How much oxygen is required to react exactly (no leftovers) with 4.3 moles of propane (C_3H_8) if the balanced equation is C_3H_8 (g) + 5 O_2 (g) $\rightarrow$ 3 CO_2 (g) + 4 H_2O(g) ?

Using the same balanced equation from above, calculate the number of moles of CO_2 formed when 4.3 mol of propane reacts with the required amount (as calculated above) of oxygen.

Section 9.3

We now know how to use a balanced chemical equation to determine mole relationships in a reaction (predict amounts of products and reactants obtained or needed).

Unfortunately, moles represent molecules and we can't count molecules. But, fortunately, we are able to "count by weighing" and can relate number of moles to a mass amount in grams using the atomic mass (elements) or molar mass (compounds).

These types of calculations are very practical and are how chemistry is actually applied in real world situations.

These calculations require the use of gram to mole and mole to gram calculations.

For Example:

If I have 35 g of aluminum, how much I_2 in grams, should I have to exactly react with the aluminum according to the balanced equation: $2Al\ (s) + 3\ I_2\ (s) \rightarrow 2AlI_3\ (s)$?

The first step is always to make sure you have a balanced reaction.

Yes, I have a balanced reaction.

Next, determine the mole relationships (mole ratios) that are given in the equation (that is write out your equivalency statements).

2 mol Al = 3 mol I_2
2 mol Al = 2 mol AlI_3
3 mol I_2 = 2 mol AlI_3

The third step is to choose the equivalency statement (mole ratio) that will allow you to convert what you need to convert.

In this question we are asked to determine amount of I_2 per amount of Al.
We need to use the statement 2 mol Al = 3 mol I_2.
This statement tells us that for every two moles of aluminum we will need 3 moles of I_2.

In order to use the equivalency statement, I must first convert grams of aluminum to moles of aluminum.

26.98 grams of aluminum = 1 mole of aluminum.

35 grams of aluminum / 26.98 grams/ 1 mole = 1.30 moles of aluminum

Now that I have moles of aluminum I can calculate moles of I_2 needed to react with 1.3 moles of aluminum using the equivalency statement 2 mol Al = 3 mol I_2.

	3 moles of I_2	
1.3 ~~moles of Al~~		= 1.3 * 3/2 = 1.95 moles of I_2
	2 ~~moles of Al~~	

I now know how many moles of I_2 I need, however my question asked how many grams of I_2 I need.

So I must now convert moles of I_2 to grams of I_2.

I need to calculate the molar mass of I_2 (2 x 126.9)

1.95 mol I_2 x 253.8 grams/ 1 mole = 495 grams I_2

There are 5 main steps to these problems

Step 1: Balance the equation

Step 2: Convert given masses of reactants or products to moles

Step 3: Using the balanced equation, set up the appropriate mole ratios (equivalency statements).

Step 4: Use the mole ratios to convert from one substance to another.

Step 5: Convert the newly calculated amount of moles back to grams.

CONGRATULATIONS!! You can now do STOICHIOMETRY.

STOICHIOMETRY is defined as the process of using a chemical equation to calculate the relative masses of reactants and products involved in a reaction.

Can be used to determine how much product can be obtained from a reaction, to compare two reactions, and lots of other great fun stuff!

Section 9.4

We now know how to use a balanced chemical equation to determine mole relationships in a reaction (predict amounts of products and reactants obtained or needed) and then use these mole relationships to determine mass relationships by converting mass to moles and moles to mass.

We now need to learn how to work with reactions when there is not enough of one reactant to completely react with another. For example:

If we react methane and water to produce hydrogen and carbon monoxide gases according to the balanced reaction: CH_4 (g) + H_2O (g) $\rightarrow$ $3H_2$ (g) + CO (g).
We can calculate that we need 279 grams of water to react exactly with 249 grams of methane.

249 grams CH_4 /16.04 g/mol= 15.5 moles of CH_4

15.5 mole of CH_4 x 1 mole of H_2O / 1 mole of CH_4= 15.5 moles of H_2O

15.5 mole of H_2O x 18.02 g/mol= 279 grams of H_2O

But what would happen if we had only 250 grams of water?

We would run out of water before we could use up all of the methane.

This then becomes a LIMITING REACTANT PROBLEM. Any problems where reactants are not mixed in exact stoichiometric quantities (for example as calculated above 249 grams to 279 grams) will require limiting reactant calculations.

A LIMITING REATANT or LIMITING REAGENT is defined as the substance that limits how much product will be formed (the substance that runs out first).

***** The most important step in a limiting reactant problem is correctly identifying the limiting reactant.

HOW TO DETERMINE THE LIMITING REACTANT

Step 1. Make certain you have a balanced equation.

Step 2. Convert masses of reactants to moles of reactants.

Step 3. Using mole ratios, compare reactants and determine which is limiting

Step 4. Using the limiting reactant, determine the amount of product(s) in moles

Step 5. Convert moles of product(s) to grams of product(s)

Let's apply this procedure to determine how many grams of hydrogen gas we would obtain if we reacted 249 grams of methane with only 250 grams of water.

Step 1. We were given the balanced equation $CH_4\ (g) + H_2O\ (g) \rightarrow 3H_2\ (g) + CO\ (g)$.

Step 2.

$$\frac{249\ \cancel{\text{grams of } CH_4}}{} \left| \frac{1 \text{ moles of } CH_4}{16.04\ \cancel{\text{grams of } CH_4}} \right. = 15.5 \text{ moles of } CH_4$$

$$\frac{250\ \cancel{\text{grams of } H_2O}}{} \left| \frac{1 \text{ moles of } H_2O}{18.02\ \cancel{\text{grams of } H_2O}} \right. = 13.9 \text{ moles of } H_2O$$

** CAUTION: Your limiting reactant is not always the substance with the fewest numbers of moles present or the fewest numbers of grams present. You need to use the mole ratios from the reaction to actually compare them.

Step 3. We must now use mole ratios (obtained from the balanced equation) to compare.

It does not mater which one you start with.

$$\frac{13.9\ \cancel{\text{moles of } H_2O}}{} \left| \frac{1 \text{ mole of } CH_4}{1\ \cancel{\text{mole of } H_2O}} \right. = 13.9 \text{ moles of } CH_4 \text{ are needed to react with } 13.9 \text{ moles of } H_2O.$$

If 13.9 moles of CH_4 are needed to react with 13.9 moles of H_2O, and we have 15.5 moles of CH_4 then there will be CH_4 left over. If the CH_4 is left over, the water will run out first, therefore water is the limiting reactant so we must calculate how much product we will obtain based off of the amount of water, not the amount of methane.

Step 4. We now calculate moles of product from our limiting reactant.

Using the limiting reactant water, calculate the number of moles of hydrogen gas that would be produced.

$$13.9\ \cancel{\text{moles of } H_2O} \times \frac{3 \text{ mole of } H_2}{1\ \cancel{\text{mole of } H_2O}} = 41.7 \text{ moles of } H_2 \text{ are produced}$$

Step 5. Convert moles of product produced to grams of product

$$41.7\ \cancel{\text{moles of } H_2} \times \frac{2.016 \text{ grams of } H_2}{1\ \cancel{\text{mole of } H_2}} = 84.07 \text{ grams of } H_2 \text{ are produced}$$

Try this one on your own:

25000 g of nitrogen gas and 5000 g of hydrogen gas are reacted to form ammonia according to the balanced equation $N_2\,(g) + 3H_2\,(g) \rightarrow 2NH_3\,(g)$. How many grams of ammonia does this create?

Step 1. We were given the balanced equation $N_2\,(g) + 3H_2\,(g) \rightarrow 2NH_3\,(g)$.
Step 2. Convert grams of reactants to moles of reactants.

$$25000\ \cancel{\text{g of } N_2} \times \frac{1 \text{ moles of } N_2}{28.02\ \cancel{\text{grams of } N_2}} = 892 \text{ moles of } N_2$$

$$5000\ \cancel{\text{g of } H_2} \times \frac{1 \text{ moles of } H_2}{2.016\ \cancel{\text{grams of } H_2}} = 2480 \text{ moles of } H_2$$

Step 3. We must now use mole ratios (obtained from the balanced equation) to compare. It does not mater which one you start with.

$$2480\ \cancel{\text{moles of } H_2} \times \frac{1 \text{ mole of } N_2}{3\ \cancel{\text{mole of } H_2}} = 826 \text{ moles of } N_2 \text{ are needed to react with } 2480 \text{ moles of } H_2.$$

If 826 moles of N_2 are needed to react with 2480 moles of H_2, and we have 892 moles of N_2 then there will be N_2 left over. If the N_2 is left over, the hydrogen gas will run out first, therefore hydrogen gas is the limiting reactant so we must calculate how much product we will obtain based off of the amount of hydrogen gas, not the amount of nitrogen gas. ** NOTE: This is an example of a case where the reactant present in the smallest actual amount of moles was not the limiting reactant due to the ratio required by the reaction.

Step 4. We now calculate moles of product from our limiting reactant.

Using the limiting reactant hydrogen gas, calculate the number of moles of ammonia that would be produced.

$$2480\ \cancel{\text{moles of } H_2} \times \frac{2 \text{ mole of } NH_3}{3\ \cancel{\text{mole of } H_2}} = 1653 \text{ moles of } NH_3 \text{ are produced}$$

Step 5. Convert moles of product produced to grams of product.

$$1653\ \cancel{\text{moles of } NH_3} \times \frac{17.03 \text{ grams of } NH_3}{1\ \cancel{\text{mole of } NH_3}} = 28150 \text{ grams of } NH_3 \text{ are produced}$$

Section 9.5

The amounts of products that we have been calculating thus far in chapter 9 are considered to be theoretical yields.

A THEORETICAL YIELD is a calculated value that assumes perfect conditions and perfect reactions.

Life is seldom perfect, chemical reactions are no different.

PERCENT YIELD is used to compare actual yield of a reaction to the calculated theoretical yield.

Actual yield is obtained in a laboratory.

Theoretical yield is calculated, just like you have been doing.

If we apply this definition to recipe scenario, the actual yield would be how many cookies you actually finished baking; the theoretical yield would be how many the recipe said you would get. They are rarely the same. (Remember you ate some of the cookie dough, you dropped one on the floor, your estimate of a rounded teaspoonful is different than someone else, you turned the mixer on too high and the dough went splattering all over the kitchen, etc. All of these things change your actual yield of cookies from the intended theoretical yield of the recipe)

If you wanted to know how efficient you were at following the recipe, you would have to take the actual number of cookies obtained divided by the number the recipe said you would get and then multiply by 100 to get a percent.

For example: The recipe said that it makes 100 cookies, you only got 78 cookies.

$$78/100 = .78 \times 100 = 78\%$$

Calculating PERCENT YIELD for a chemical reaction is no different.

Percent yield is calculated by taking actual yield and dividing it by the theoretical yield and then multiplying by 100 to obtain a percent.

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

Let’s use the ammonia problem from the last section to explore this type of calculation. We already calculated the theoretical yield of ammonia from the reaction of 25000 g of nitrogen gas and 5000 g of hydrogen gas to be 28150 grams of NH_3. But what if the reaction actually yielded only 27000 grams of NH_3? What would the percent yield of the reaction be in this scenario?

Just to review, here are the calculations for the theoretical yield.

Step 1. We were given the balanced equation $N_2\,(g) + 3H_2\,(g) \rightarrow 2NH_3\,(g)$.

Step 2. Convert grams of reactants to moles of reactants.

$$25000\ \cancel{\text{g of } N_2} \times \frac{1 \text{ moles of } N_2}{28.02\ \cancel{\text{grams of } N_2}} = 892 \text{ moles of } N_2$$

$$5000\ \cancel{\text{g of } H_2} \times \frac{1 \text{ moles of } H_2}{2.016\ \cancel{\text{grams of } H_2}} = 2480 \text{ moles of } H_2$$

Step 3. We must now use mole ratios (obtained from the balanced equation) to compare. It does not mater which one you start with.

$$2480\ \cancel{\text{moles of } H_2} \times \frac{1 \text{ mole of } N_2}{3\ \cancel{\text{mole of } H_2}} = 826 \text{ moles of } N_2 \text{ are needed to react with } 2480 \text{ moles of } H_2.$$

If 826 moles of N_2 are needed to react with 2480 moles of H_2, and we have 892 moles of N_2 then there will be N_2 left over. If the N_2 is left over, the hydrogen gas will run out first, therefore hydrogen gas is the limiting reactant so we must calculate how much product we will obtain based off of the amount of hydrogen gas, not the amount of nitrogen gas. ** NOTE: This is an example of a case where the reactant present in the smallest actual amount of moles was not the limiting reactant due to the ratio required by the reaction.

Step 4. We now calculate moles of product from our limiting reactant.

Using the limiting reactant hydrogen gas, calculate the number of moles of ammonia that would be produced.

$$\frac{2480\ \cancel{\text{moles of } H_2}}{} \left| \frac{2 \text{ mole of } NH_3}{3\ \cancel{\text{mole of } H_2}} \right. = 1653 \text{ moles of } NH_3 \text{ are produced}$$

Step 5. Convert moles of product produced to grams of product.

$$\frac{1653\ \cancel{\text{moles of } NH_3}}{} \left| \frac{17.03 \text{ grams of } NH_3}{1\ \cancel{\text{mole of } NH_3}} \right. = 28150 \text{ grams of } NH_3 \text{ are produced}$$

Once we have calculated the theoretical yield we can compare it to the actual yield. I will have to give you this in the problem.

If the actual yield of the reaction is 27000 grams of NH_3, what is the percent yield of the reaction?

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

$$\boxed{\text{Percent Yield} = \frac{27000 \text{ grams of } NH_3}{28150 \text{ grams of } NH_3} \times 100}$$

Percent Yield = 95.9%

Just as with percent composition calculations, the math involved with these calculations is just like the math you use to determine what percent you achieved on a quiz or test.

This is not the only application of percent yield problems.

You can also be given percent yield and be asked to calculate actual yield.

For Example: If I told you that the percent yield for the same reaction, under different circumstances was only 73%, how many grams of ammonia were obtained in this scenario. This just required an adaptation of the same formula.

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

$$73 = \frac{\text{Actual grams of } NH_3}{28150 \text{ grams of } NH_3} \times 100$$

Algebraically rearrange the formula and calculate the actual yield of ammonia

$$\underline{\text{????? grams of } NH_3} = \frac{73 \times 28150 \text{ grams of } NH_3}{100}$$

The actual yield (in grams) of ammonia would be 20549 grams.