CHEMISTRY NOTES – Chapter 9

Stoichiometry

Goals: To gain an understanding of:

1. Stoichiometry.
2. Limiting reagents and percent yield.

NOTES:

Stoichiometry is the calculation of chemical quantities from balanced equations.

The four quantities involved in stoichiometric calculations are:

- particles - the relative amounts of atoms, ions, unit formulas or molecules in various reactants or products
- moles - the relative number of moles of reactants or products
- mass - the relative masses of reactants or products
- volume - the relative amounts of gaseous reactants or products

Atoms and mass are always conserved in chemical reactions.

Before getting into stoichiometric calculations let me make an analogy to an everyday experience: the making of a peanut butter and jelly sandwich. Here is the equation:

2 slices bread + 1 serving peanut butter (pb) + 1 serving jelly → 1 sandwich

In this "balanced equation" we see a number of ratios: 2 slices bread/1 serving pb, 2 slices bread/1 serving jelly or 1 serving pb/1 serving jelly etc..

Stoichiometry uses these ratios to determine relative amounts of reactants or products. For example, if I had 12 slices of bread, how many servings of peanut butter would I need and how many sandwiches could I make? Here are the stoichiometric calculations:

Here is an example of a balanced chemical equation with some chemical quantities listed below them.

\[
4 \text{ Na}_2\text{O(s)} + O_2(g) \rightarrow 2 \text{ Na}_2\text{O}_2(s)
\]

Here are some calculations that can then be performed:

a. If 3.50 moles of Na₂O are needed, how many moles of sodium will be required (assuming sufficient oxygen)?

\[
3.50 \text{ moles Na}_2\text{O} \times \frac{4 \text{ moles Na}}{2 \text{ moles Na}_2\text{O}} = 7.00 \text{ moles Na}
\]
b. If you have 45.00 g Na and an abundant supply of oxygen, how much Na₂O should be produced?

$$\text{45.00 g Na} \left( \frac{123.96 \text{ g Na}_2\text{O}}{91.96 \text{ g Na}} \right) = 60.66 \text{ g Na}_2\text{O}$$

c. If you had an abundant supply of oxygen and needed to produce 12.2 moles of Na₂O, how many grams of sodium would you need?

$$\frac{12.2 \text{ mol Na}_2\text{O}}{2 \text{ mol Na}_2\text{O}} \left( \frac{4 \text{ mol Na}}{1 \text{ mol Na}_2\text{O}} \right) \left( \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} \right) = 561 \text{ g Na}$$

d. How many formula units of Na₂O could be produced from 500.0 g of sodium metal? This one is a little trickier so you need to plan ahead. From grams of Na you can determine moles of Na. From moles Na you can calculate moles Na₂O from the mole ratio, and then from moles of Na₂O you can calculate unit formulas of Na₂O.

grams Na → moles Na → moles Na₂O → formula units or molecules

$$\frac{500.0 \text{ g Na}}{22.99 \text{ g Na/mol}} \left( \frac{4 \text{ mol Na}_2\text{O}}{1 \text{ mol Na}} \right) \left( \frac{6.02 \times 10^{23} \text{ formula units}}{1 \text{ mol Na}_2\text{O}} \right) = 2.61 \times 10^{25} \text{ formula units}$$

The limiting reagent is the reactant which is in the smallest proportional amount and so determines the amount of product that will be produced. For example if we had 10 servings of both peanut butter and jelly, but only 4 slices of bread we could only make two complete peanut butter and jelly sandwiches. In this example the four slices of bread is the limiting reagent and so determines how many sandwiches (product) could be produced. The servings of peanut butter and jelly could be described as excess reagents.

Let's try a real chemistry problem and use this concept to determine the limiting reagent and see how it determines how much product will be formed.

Problem: Nitrogen gas reacts with hydrogen gas to produce gaseous ammonia (NH₃). If we react 25.0 g of nitrogen with 25.0 g of hydrogen, how much ammonia will be produced. (Identify the limiting and excess reagents and determine how much of the excess reagent will remain).

Step 1. Write the balanced equation: \(\text{N}_2(g) + 3 \text{H}_2(g) \rightarrow 2 \text{NH}_3(g)\)

Step 2. Determine the limiting reagent. Do this by first converting reactant quantities to moles and then using the mole ratio to convert from one reactant to the other (your choice).

Converting mass to moles:
- \(25.0 \text{ g H}_2 \left( \frac{1 \text{ mol H}_2}{2.02 \text{ g}} \right) = 12.4 \text{ mol H}_2\)
- \(25.0 \text{ g N}_2 \left( \frac{1 \text{ mol N}_2}{28.02 \text{ g}} \right) = 0.892 \text{ mol N}_2\)

Converting mol H₂ to mol N₂:
- \(12.4 \text{ mol H}_2 \left( \frac{1 \text{ mol N}_2}{3 \text{ mol H}_2} \right) = 4.13 \text{ mol N}_2\)
From above we can see that if we have 12.4 mol H\textsubscript{2} we need 4.13 mol N\textsubscript{2}. We don't have that much N\textsubscript{2} so the .892 mol of N\textsubscript{2} must be the limiting reagent.

We can now determine how much ammonia will be produced using the mole ratio in the balanced equation:

\[
.892 \text{ mol } N_2 \left( \frac{2 \text{ mol NH}_3}{1 \text{ mol } N_2} \right) = 1.78 \text{ mol NH}_3
\]

We can also determine how much H\textsubscript{2} (excess reagent) will remain by determining how many moles of hydrogen will be used and then convert to grams and subtract from the excess amount given in the problem.

### Determining moles of hydrogen used:

\[
.892 \text{ mol } N_2 \left( \frac{3 \text{ mol } H_2}{1 \text{ mol } N_2} \right) = 2.68 \text{ mol } H_2
\]

### Converting moles to mass:

\[
2.68 \text{ mol } H_2 \left( \frac{2.02 \text{ g } H_2}{1 \text{ mol } H_2} \right) = 5.41 \text{ g } H_2
\]

### Determining amount of excess:

\[
25.0 \text{ g } - 5.41 \text{ g } = 19.6 \text{ g } H_2
\]

The theoretical yield is the amount of product you should get based on the calculations as done above (e.g. the 1.78 moles of NH\textsubscript{3}). The theoretical yield is what you get under perfect conditions where 100% of the limiting reagent reacts to produce product.

The actual yield is the amount you actually do get under the industrial or laboratory conditions of the reaction.

The percent yield is calculated by the equation: (actual yield/theoretical yield) x 100%. For example, let's say we actually did the above experiment - took 25.0 g of nitrogen and reacted it with 25.0 g of hydrogen and produced 29.6 g of ammonia (actual yield). How would we calculate the percent yield? The first step is to get common units (your choice) - the actual yield is in grams and the theoretical yield above is in units of moles. I will convert the theoretical yield to grams and calculate the percent yield as shown below.

### Converting theoretical yield into grams:

\[
1.78 \text{ mol } NH_3 \left( \frac{17.04 \text{ g}}{1 \text{ mol}} \right) = 30.03 \text{ g}
\]

### Calculating percent yield:

\[
\% \text{ yield} = \left( \frac{29.6 \text{ g}}{30.3 \text{ g}} \right) \times 100\% = 97.7\%
\]