Lecture 2 Notes, Immaculata Week, July-August 2014, Charles H. Mahler, Lycoming College

My email address is mahler@lycoming.edu – please contact me if you have questions.

Our topic for this lecture was further uses of the mole concept, especially the ideas of limiting reagent and theoretical and percent yield.

We started with a problem to review our skills from the previous session. In 2008, power plants in the US burned 1.14 x 10^{11} kg of natural gas to generate electricity. Assuming the natural gas was all methane, how many kg of carbon dioxide and water were released by these power plants?

The final balanced reaction burning methane is: 1 CH_4 + 2 O_2 → 1 CO_2 + 2 H_2O. Rounded to the ones place, methane s 16 g/mol, oxygen gas is 32 g/mol, carbon dioxide is 44 g/mol and water is 18 g/mol.

(1.14 x 10^{11} kg)(1000 g/kg)(1 mole methane/16 g) = 7.124 x 10^{12} moles methane burnt

Since the stoichiometry between methane and carbon dioxide is 1:1, the same number of moles of carbon dioxide is produced.

(7.124 x 10^{12} moles CO_2)(44 g/mol)(1 kg/1000g) = 3.13 x 10^{11} kg CO_2 produced

For water produced the stoichiometry is 1:2 (for every one mole of methane that burns, two moles of water is produced. So the moles of water are twice the moles of methane (and of carbon dioxide), or 1.425 x 10^{13} moles H_2O.

(1.425 x 10^{13} moles H_2O)(18.02 g/mol)(1 kg/1000 g) = 2.57 x 10^{11} kg H_2O produced

We discussed how the answers need to make sense – we start with roughly 10^{11} kg methane and get roughly 10^{11} kg carbon dioxide and water (with more of each from increased molar mass and the 1:2 water stoichiometry). It is always good to look at the answer and see if it seems reasonable and makes sense. We also noted that even methane, which is the greenest fossil fuel (since it produces the least carbon dioxide) still produces a staggering amount of the greenhouse gas carbon dioxide when burned in the quantities used in power plants nationwide.

In some classes we talked about weight percentages and finding formulas from them. For example a compound that contains only iron, Fe, and oxygen, O, has 77.23% Fe and 22.27% O. What is its formula? We start by assuming there are 100 grams of the compound, so there are 77.23 g of Fe and 22.27 g of O. The moles of iron are found by dividing the mass of iron by the atomic weight of iron: 77.73g Fe/55.85 g/mol = 1.392 moles Fe

For the moles of oxygen: 22.27 g O/16.00 g/mol = 1.392 moles O

We look at the simplest ratio for the empirical formula, so 1:392 to 1.392 is 1:1 or “FeO”.

We also looked at a compound that was 69.94% Fe and 20.06% O. Using the same approach:

69.94 g Fe/55.85 g/mol = 1.252 mol Fe and 30.06 g O/16.00 g/mol = 1.879 mol O
Dividing the larger number of moles by the smaller number of moles 1.879/1.252 = 1.500, so there are 1.5 moles oxygen for every 1 mole of iron. Since formulas have to be integers, we multiply by 2 to find that there are 3 moles of oxygen for every 2 moles of iron, or Fe_2O_3

Next we turned to the idea of limiting reagent or limiting reactant. This is a concept we already know as “what runs out first”? If you make a cheese sandwich, you need 2 slices of bread for every 1 slice of cheese, so if you have 5 slices of cheese and 8 slices of bread, you can only make 4 sandwiches (since the bread runs out first). You will have leftover cheese (even though there were fewer slices of cheese than bread). You can only make sandwiches until bread runs out, it is the limiting reactant in this case. Next we move on to a chemistry example.

Consider the reaction: $SO_3 + H_2O \rightarrow H_2SO_4$ If you start with 20.00 grams of SO_3 and 10.00 grams of H_2O, which is the limiting reactant (which runs out first)? How much of the product (sulfuric acid, H_2SO_4) will be produced?

We need to find the moles of each reactant, and then the moles (or grams) or product each could produce.

(20.00 g SO_3)(1 mol/80.06 g) = 0.2498 moles SO_3 = 0.2498 moles or 24.50 g H_2SO_4 (as 1:1 stoichiometry)

(10.00 g H_2O)(1 mol/18.02 g) = 0.5549 moles H_2O = 0.5549 moles or 54.42 g H_2SO_4 (as 1:1 stoichiometry)

Since there are fewer moles of sulfuric acid, H_2SO_4, produced from SO_3, then SO_3 is the limiting reactant. You use up all of the SO_3 and make no more than 0.2498 moles H_2SO_4 (which is 24.50 grams). You cannot make more sulfuric acid than this, because the SO_3 has been completely used up or consumed.

You can also figure out how much water is left over (remains unreacted). Since we start with 0.5549 moles of water and 0.2498 moles of sulfuric acid are made, then because of the 1:1 stoichiometry between sulfuric acid and water, 0.2498 moles of water are consumed. What is left in terms of moles of water is 0.5549 mol – 0.2498 mol = 0.3051 mole of water remain unreacted. You can multiply this times the molar mass of water, 18.02 g/mol to find this is 5.50 grams of water left.

Since there are only two reactants and one product, you could also look at conservation of mass to find the mass of water left unreacted. There are 24.50 grams of product made (H_2SO_4). Since SO_3 is the limiting reactant, it is all gone and 20.00 grams of product come from it. That means the remaining 4.50 grams come from the water. Since we started with 10.00 grams of water, what is left over is 10.00 – 4.50 = 5.50 grams of water left.

The last topic we looked at were various kinds of yields. The amount of product expected from the limiting reactant is the theoretical yield. However, often after a reaction is worked up there is not as much product as expected. Sometimes this is due to losses in processing (small amounts stick to glassware or are lost while being transferred) or sometimes this is due to competing reactions that produce small amounts of other products. For whatever reason, the experimental yield or actual yield is often less.

You can look at the ratio of the actual yield to the theoretical yield to find the percent yield:

$$\frac{\text{Actual yield x 100\%}}{\text{Theoretical yield}} \text{ - percent yield}$$

So if our reaction expected to produce 24.50 grams above (the theoretical yield) actually resulted in an experimental amount of product of 20.79 grams, then the percent yield would be:

(20.79 g / 24.50 g)x 100% = 84.86% percent yield, since the product isolated was only 84.86% of what was expected.

We did a hands on experiment on Chemistry in a Ziploc bag and I demonstrated the Exploding Pringles Can (hydrogen plus oxygen makes water). Please see the handout on activities and demonstrations for more details.