6.5: Limiting Reactants

Gloves will typically come is right- and left-handed models. In order to make a *pair* of gloves, you need one that is designed to fit each hand. If you had a box containing 50 left-handed gloves and another box containing 40 right-handed gloves, you could make 40 proper pairs and you would have ten left-handed gloves left over. The number of pairs of gloves that you could assemble is *limited* by the glove in the smallest number (the right-handed glove). The other glove in this example, the left, is present in *excess.*

The same sort of logic applies to chemical reactions in which there are two or more reactants. In Example 6.4, we carefully weighed out 0.010 mole of solid zinc and solid sulfur in order to react them to form 0.010 mole of the product, ZnS. If instead, we had reacted 0.010 mole of Zn with 0.020 moles of sulfur, how much ZnS would have (theoretically) formed? The answer is still 0.010 mole of ZnS. What happens to the leftover sulfur? It just sits there! When the reaction is complete, there is (theoretically) 0.010 mole of ZnS mixed with the remaining 0.010 mole of sulfur. 10 atoms of Zn reacting with 20 atoms of S yield 10 molecules of ZnS with 10 atoms of S remaining.

You may have noticed that, in many of the problems in this chapter, we stated that one reactant reacted with an *excess* of a second reactant. In all of these cases, the theoretical yield of product is determined by the limiting reactant in the reaction, and some of the excess reactant is left over. If aluminum and chlorine gas, a diatomic gas, react to form aluminum chloride according to the equation shown below,

$$\mathbf{2\ Al}\ \textit{(s)} + \mathbf{3\ Cl_2}\ \textit{(g)} \rightarrow \mathbf{2\ AlCl_3}\ \textit{(s)}$$

and there are 2.0 moles of aluminum and 14 moles of chlorine present, two moles of aluminum chloride are formed and 11 moles of chlorine gas remain in excess. Our stoichiometry is:

$$\left(\frac{3\ mol\ Cl_2}{2\ mol\ AlCl_3}\right)$$

If three moles of chlorine gas are used in the reaction, (14 – 3) = 11 moles of chlorine must remain. When a problem is presented and one reactant is labeled as excess, the theoretical yield of product is equal to the moles of the limiting reagent, adjusted for the stoichiometry of the reaction.

Although it would be easier if reactants were routinely labeled as “limiting” or “excess”, more commonly problems are written in such a way that it is not always trivial to identify the limiting reactant. For example, if you were told that 6.0 grams of aluminum was reacted with 3.8 grams of chlorine gas and you were asked to calculate the mass of $AlCl_3$ that would be formed, there would be no simple way to identify the limiting- and excess reactants. In a case like this what you want to do is to simply solve the problem *twice*. First you would calculate the number of moles of aluminum in 6.0 grams, and then calculate how many moles of $AlCl_3$ could be formed. Next, you calculate how many moles of chlorine are present in 3.8 grams of chlorine gas and again, calculate how many moles of $AlCl_3$ could be formed. Whichever reagent produces the *smallest number of moles of product* must be *limiting* and the other reagent must be in *excess*.

Find: *x* moles of $AlCl_3$

We set up the problem to solve for *mol product* for *each* reactant. The general equation is:

$$(molproduct) = (molreactant) \times \left(\frac{molproduct}{molreactant}\right)$$

The solutions for both reactants are:

$$x\ mol\ AlCl_3 = (6.0\ g\ Al)\left(\frac{1\ mol\ Al}{26.98\ g\ Al}\right)\left(\frac{1\ mol\ AlCl}{1\ mol\ Al}\right)$$

$$x\ mol\ AlCl_3 = (3.8\ g\ Cl_2)\left(\frac{1\ mol\ Cl_2}{70.90\ g\ Cl_2}\right)\left(\frac{2\ mol\ AlCl}{3\ mol\ Cl_2}\right)$$

Solving these equations, we see that, beginning with 6.0 grams of aluminum, **0.22 moles** of $AlCl_3$ can be formed, and that, beginning with 3.8 grams of chlorine, **0.036 moles** of $AlCl_3$ can be formed.

Keeping score, 6.0 grams of Al yields 0.22 moles of $AlCl_3$, and 3.8 grams of chlorine gas yields 0.036 moles of $AlCl_3$. The *lowest* yield comes from the chlorine gas, therefore it must be *limiting* and aluminum must be in excess. The reaction in the problem will therefore produce 0.036 moles of product, which is equivalent to:

$$0.036\ mol\ AlCl_3(133.33\ g/mol) = 4.8\ g\ AlCl_3$$

Exercise 6.5.1

Lead (IV) chloride reacts with fluorine gas to give lead (IV) fluoride and Cl_2. If 0.023 moles of fluorine gas reacts with 5.3 grams of lead (IV) chloride, what mass of lead (IV) fluoride will be formed?

Although "limiting reactant problems" may be *tedious*, they are not difficult. When you are faced with a limiting reactant problem, just remember, you do the simple molar yield calculations *twice*, one for each reactant. The reactant that yields the *lowest* molar quantity is your limiting reagent and the molar value you calculate determines the theoretical yield in the problem.

Exercise 6.5.2

Ammonia, which is the active ingredient in "smelling salts", is prepared from nitrogen and hydrogen according to the equation shown below.

$$\mathbf{N_2 + 3\ H_2 \rightarrow 2\ NH_3}$$

a. If you mix 5.0 mol of nitrogen and 10.0 moles of hydrogen how many moles of ammonia would you produce? Which reactant is in excess?
b. If you have 6.2 grams of nitrogen and you react it with 6.2 grams of hydrogen how many grams of ammonia would you produce? Which reactant is the limiting reactant?