Limiting Reactants

You drop a piece of zinc into a beaker of hydrochloric acid. It begins to bubble furiously, but eventually it stops. You drop another piece of zinc into the acid. The bubbling begins anew, but again it stops. This time, you add more hydrochloric acid. Nothing happens. Obviously, each time the reaction stopped it was because you ran out of zinc. There was always plenty of hydrochloric acid. In fact, there was an excess of the acid. Zinc, on the other hand, was a limiting reactant. The reactant that is consumed first limits the amount of product that is produced, and is called a limiting reactant. The balanced equation predicts the amounts of reactants needed to completely consume each other (stoichiometric quantities). If any of the reactants is in excess, the other(s) is (are) limiting reactants.

Limiting reactants in the *real* world.

In stoichiometry problems where the amount of only one of the reactants is given, it is assumed either that there are stoichiometric amounts of all the reactants and products, or, at the very least, that all of the reactant for which the amount is specified is consumed because it is the limiting reactant. For problems in which the amount of more than one reactant is specified, you need to consider that there may be a limiting reactant. If there is a limiting reactant, you need to know which one it is and use it for your calculations, because excess, unconsumed reactants do NOT produce any product.

To identify the limiting reactant: [1] Write a balanced equation; [2] Calculate the number of moles of each of the reactants present; and [3] Divide the number of moles of each reactant by its stoichiometric coefficient. The smallest number corresponds to the limiting reactant. Once the limiting reactant is identified, stoichiometry problems are done as usual using the amount of the limiting reactant.

Sample Problem

How much ammonia is formed from 25.0 kg of nitrogen and 5.00 kg of hydrogen? How much material is unreacted?

Step 1: Write the balanced equation

$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

Step 2: Calculate the number of moles of each of the reactants present.

$$\left(25.0kg_{N_2}\right)\left(\frac{1000g}{1kg}\right)\left(\frac{1mol_{N_2}}{28.0g_{N_2}}\right) = 8.93\times10^2 mol_{N_2} \qquad \left(5.00kg_{H_2}\right)\left(\frac{1000g}{1kg}\right)\left(\frac{1mol_{H_2}}{2.02g_{H_2}}\right) = 2.48\times10^3 mol_{H_2}$$

Step 3: Divide the number of moles of each reactant by its stoichiometric coefficient. The smallest number corresponds to the limiting reactant.

$$\frac{8.93\times10^2 mol_{N_2}}{1} = 8.93\times10^2 mol_{N_2} \qquad \frac{2.48\times10^3 mol_{H_2}}{3} = 8.25\times10^2 mol_{H_2}$$

H_2 is limiting

Step 4: Use the limiting reactant to complete the calculation.

$$\left(2.48\times10^3 mol_{H_2}\right)\left(\frac{2mol_{NH_3}}{3mol_{H_2}}\right)\left(\frac{17.0g_{NH_3}}{1mol_{NH_3}}\right)\left(\frac{1kg}{1000g}\right) = 28.1kg_{NH_3}$$

Step 5: Calculate the amount of unreacted material by calculating the stoichiometric amount that reacted and subtracting it from the initial amount.

$$\left(2.48\times10^3 mol_{H_2}\right)\left(\frac{1mol_{N_2}}{3mol_{H_2}}\right)\left(\frac{28.0g_{N_2}}{1mol_{N_2}}\right)\left(\frac{1kg}{1000g}\right) = 23.1kg_{N_2}$$

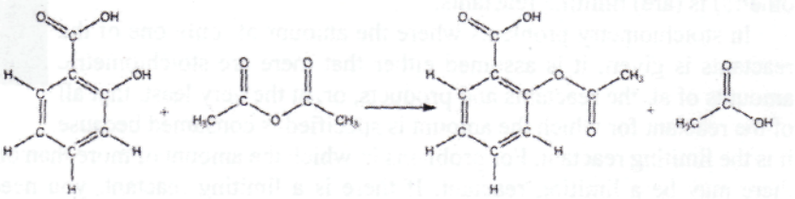
$$25.0kg_{N_2} - 23.1kg_{N_2} = 1.9kg_{N_2}$$

(*CONTINUED ON THE NEXT PAGE*)

Answer the questions below based on the preceding example. (*NOTE*: Equations provided may not be balanced.)

1. How much pure iron can be extracted from 250. kg of iron III oxide when it reacts with 148 kg of carbon monoxide? What is in excess, and by how much? [$Fe_2O_3 + CO \rightarrow Fe + CO_2$]

2. Salicylic acid ($C_7H_6O_3$) reacts with acetic anhydride ($C_4H_6O_3$) to form aspirin ($C_9H_8O_4$) and acetic acid ($C_2H_4O_2$). How much aspirin forms from 25.0 g of salicylic acid and 25.0 g of acetic anhydride? What is in excess, and by how much?
   [$C_7H_6O_3 + C_4H_6O_3 \rightarrow C_9H_8O_4 + C_2H_4O_2$].



3. How much copper will precipitate when 10.0 g of granular zinc are added to solution containing 16.0 g of aqueous copper II sulfate? What is in excess, and by how much? [$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$]

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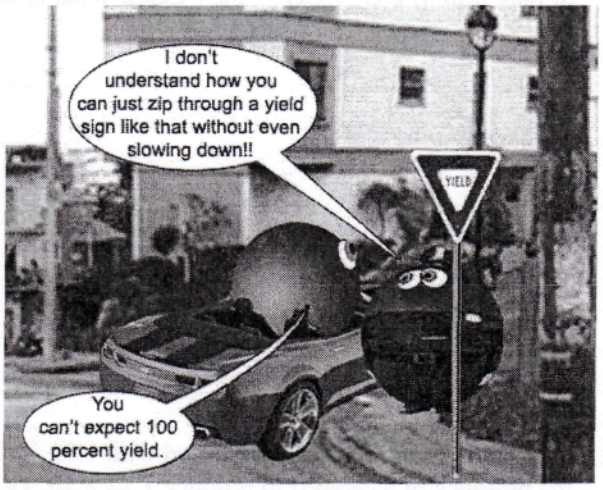
Percent Yield

You are making chocolate chip cookies. The recipe calls for two eggs to make 2½ dozen cookies. You have only one egg. There is no choice but to make half a recipe. The next problem is that your oven doesn't heat evenly, so the cookies towards the back are always better done than the ones in front. You pop them into the oven and hope for the best. In the end, the three cookies in the back row are burnt beyond recognition, but the rest are good. By scaling back the recipe, you could have theoretically anticipated having 15 cookies, but you only have 12 that are edible. Your yield is only 80 percent of what you anticipated. The same thing happens in chemistry.

Chocolate Chip Cookies
(makes 2½ dozen cookies)

3 cups flour
1¼ teaspoons salt
1 teaspoon baking soda
¼ teaspoon baking powder
¾ cup unsalted butter
1 cup dark brown sugar
½ cup white sugar
1 tablespoon vanilla extract
2 eggs
2 tablespoons corn syrup
1 tablespoon half-and-half
2 cups chocolate chips



Why atoms make notoriously poor drivers.

In chemical reactions, the actual yield is usually less than the theoretical yield due to side reactions and other complications. The theoretical yield is the amount of product formed when the limiting reactant is completely consumed. It is the maximum amount of product that can be produced. The actual yield is often expressed as a percentage of the theoretical yield called the percent yield.

$$percent\ yield = \frac{actual\ yield}{theoretical\ yield} \times 100$$

Sample Problem

If 68.5 kg of CO reacts with 8.60 kg of H_2 to produce 35.7 kg of CH_3OH, what is the percent yield?

Step 1: Write a balanced equation.

$2H_2(g) + CO(g) \rightarrow CH_3OH(\ell)$

Step 2: Identify the limiting reactant

$$(8.60kg_{H_2})\left(\frac{1000g}{1kg}\right)\left(\frac{1mol_{H_2}}{2.02g_{H_2}}\right) = 4.26\times10^3 mol_{H_2} \qquad \frac{4.26\times10^3 mol_{H_2}}{2} = 2.13\times10^3 mol_{H_2}$$

$$(68.5kg_{CO})\left(\frac{1000g}{1kg}\right)\left(\frac{1mol_{CO}}{28.0g_{CO}}\right) = 2.45\times10^3 mol_{CO} \qquad \frac{2.45\times10^3 mol_{CO}}{1} = 2.45\times10^3 mol_{CO}$$

H_2 *is limiting*

Step 3: Use the limiting reactant to complete the calculation of the theoretical yield

$$(4.26\times10^3 mol_{H_2})\left(\frac{1mol_{CH_3OH}}{2mol_{H_2}}\right)\left(\frac{32.04g_{CH_3OH}}{1mol_{CH_3OH}}\right)\left(\frac{1kg}{1000g}\right) = 68.2kg_{CH_3OH}$$

Step 4: Calculate the percent yield

$$\frac{35.7kg_{CH_3OH}}{68.2kg_{CH_3OH}} \times 100 = 52.3\%$$

(*CONTINUED ON THE NEXT PAGE*)

Calculate the percent yield in each of the problems below. (*NOTE*: Equations may not be balanced.)

1. 3.25×10^3 kg of iron III oxide are treated in a blast furnace with 1.50×10^3 kg of carbon monoxide to form 1.30×10^3 kg of pure iron. [$Fe_2O_3 + CO \rightarrow Fe + CO_2$]. What is the percent yield?

2. 6.16×10^2 kg of nitrogen are reacted with 1.75×10^2 kg hydrogen under conditions of high temperature and pressure in the presence of a catalyst to produce 1.12×10^2 kg ammonia. [$N_2 + H_2 \rightarrow NH_3$] What is the percent yield?

3. The explosive, ammonium nitrate, is produced by reacting ammonia with nitric acid (HNO_3). [$NH_3 + HNO_3 \rightarrow NH_4NO_3$] If 475 kg of ammonia are reacted with 1,060 kg of nitric acid to produce 1150 kg of ammonium nitrate, what is the percent yield?

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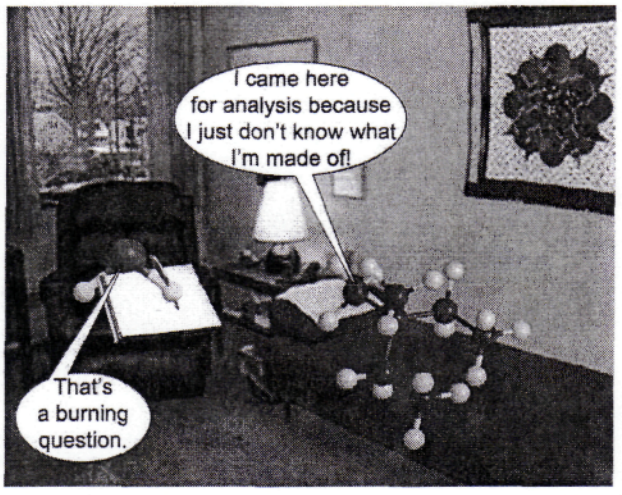
Chemistry: Form WS5.14.1A Name _____

TOPIC Date _____ Period _____

Analysis by Combustion

You have an unknown hydrocarbon. What is it? One way to analyze it is by burning it. Hydrocarbons burn to produce two oxides, water and carbon dioxide. It is possible to measure the masses of the oxides formed and determine the nature of the compound they came from. This works because you know the percent composition of both the water and the carbon dioxide. As a result, it is possible to find the mass of carbon and the mass of hydrogen in the original compound. From these masses, it is possible to determine the percent composition of the original compound and the empirical formula.



Sample Problem

A pure hydrocarbon is burned in an excess of oxygen and produces 616.6 g of carbon dioxide and 283.5 g of water. What is it's empirical formula?

Step 1: Determine the number of grams of carbon and the number of grams of hydrogen burned based on the percent composition of their oxides.

a) Find the formula masses of each

CO_2	H_2O
C = 12.01 × 1 = 12.01	H = 1.008 × 2 = 2.016
O = 16.00 × 2 = 32.00	O = 16.00 × 2 = 16.00
44.01	18.02

b) Find the percentages of each.

$\%C: \frac{12.01}{44.01} \times 100 = 27.29\%$ $\%H: \frac{2.016}{18.02} \times 100 = 11.19\%$

c) determine the number of grams of each

C: 616.6 g × 0.2729 = 168.3 g H: 283.5 g × 0.1119 = 31.72 g

Step 2: Determine the mole ratios

	C	H
moles	$\frac{168.3}{12.01} = 14.01$	$\frac{31.72}{1.008} = 31.47$
ratio	$\frac{14.01}{14.01} = 1$	$\frac{31.47}{14.01} = 2.25$
integer	$1 \times 4 = 4$	$2.25 \times 4 = 9$
formula	C_4H_9	

Answer the questions below as shown in the preceding example.

1. A pure hydrocarbon is burned in an excess of oxygen and produces 1,813 g of carbon dioxide and 990.2 g of water. What is it's empirical formula?

2. A pure hydrocarbon is burned in an excess of oxygen and produces 929.7 g of carbon dioxide and 380.6 g of water. What is it's empirical formula?

3. A pure hydrocarbon is burned in an excess of oxygen and produces 3.811 kg of carbon dioxide and 3.747 kg of water. What is it's empirical formula?

4. A pure hydrocarbon is burned in an excess of oxygen and produces 1.318 kg of carbon dioxide and 1.079 kg of water. What is it's empirical formula?

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