Answer the following questions about the reaction $CH_2Cl_2 + 2\ Cl_2$ ---> $CCl_4 + 2\ HCl$. Show your work.

1. a. If one molecule of CH_2Cl_2 reacts this way, and the reaction works perfectly, how many molecules of HCl will be formed?

   b. If one mole of Cl_2 reacts this way, and the reaction works perfectly, how many moles of CCl_4 will be formed? That is, what will be the theoretical yield of CCl_4 in moles?

   c. If one mole of CH_2Cl_2 and four moles of Cl_2 are mixed and the reaction takes place, what is the theoretical yield of CCl_4 in moles?

   d. If 1.2 moles of CH_2Cl_2 and 2.1 moles of Cl_2 are mixed, what reagent will you run out of first? That is, what is the limiting reagent?

2. a. If the instructions say to use 1.2 moles of CH_2Cl_2, how many grams of CH_2Cl_2 should you measure out?

   b) You isolate 17.8 grams of CCl_4 from a reaction. How many moles of CCl_4 is this?

3. a. If 1.0 gram of CH_2Cl_2 reacts with an excess of Cl_2, what is the theoretical yield of HCl in moles?

   b. If 10.2 grams of Cl_2 react with an excess of CH_2Cl_2, what is the theoretical yield of CCl_4 in moles?

4. a. If 2.7 grams of CH_2Cl_2 reacts with an excess of (that is, more than enough) Cl_2 this way, what is the theoretical yield of CCl_4 in grams?

   b. If you mix 8.9 g of CH_2Cl_2 with 11.1 g of Cl_2, what is the limiting reagent?

   c. If you mix 8.9 g of CH_2Cl_2 with 11.1 g of Cl_2, what is the theoretical yield of CCl_4 in grams?

5. a. If you mix 8.9 g of CH_2Cl_2 with 27.1 g of Cl_2, what is the limiting reagent?

   b. If you mix 8.9 g of CH_2Cl_2 with 27.1 g of Cl_2, what is the theoretical yield of CCl_4 in grams?

   c. If you mix 8.9 g of CH_2Cl_2 with 27.1 g of Cl_2, and get 13 g of CCl_4, what is the percent yield of CCl_4 in grams?

SOLUTIONS TO THE SAMPLE PROBLEMS:

1. a. If one molecule of CH_2Cl_2 reacts this way, and the reaction works perfectly, how many molecules of HCl will be formed?

   The balanced equation is $CH_2Cl_2 + 2\,Cl_2 \longrightarrow CCl_4 + 2\,HCl$, so 2 molecules of HCl are produced from each molecule of CH_2Cl_2. If you do this in equation form, you get:

$$\frac{1\ \cancel{\text{molecule } CH_2Cl_2}}{} \ \frac{(2\text{ molecules HCl})}{(1\ \cancel{\text{molecule } CH_2Cl_2})} = 2\text{ molecules HCl}$$

   Notice how the units cancel out. You make the conversion factor in parentheses from the coefficients in the balanced equation. This looks overly elaborate now, but will be important later.

   b. If one mole of Cl_2 reacts this way, and the reaction works perfectly, how many moles of CCl_4 will be formed? That is, what will be the theoretical yield of CCl_4 in moles?

   According to the balanced equation, two moles of Cl_2 give one mole of CCl_4. One mole of Cl_2 will therefore give half a mole of CCl_4. In equation form, this is:

$$\frac{1\text{ mole } Cl_2}{} \ \frac{(1\text{ mole } CCl_4)}{(2\text{ moles } Cl_2)} = 0.5\text{ mole HCl}$$

   c. If one mole of CH_2Cl_2 and four moles of Cl_2 are mixed and the reaction takes place, what is the theoretical yield of CCl_4 in moles?

   You didn't say five moles, did you? One mole of CH_2Cl_2 reacts with two moles of the Cl_2, leaving two of the moles of Cl_2 left over. When this happens, we say that the Cl_2 was *in excess.* One mole of CH_2Cl_2 gives 1 mole of CCl_4.

$$\frac{1\text{ mole } CH_2Cl_2}{} \ \frac{(1\text{ mole } CCl_4)}{(1\text{ mole } CH_2Cl_2)} = 1\text{ mole } CCl_4$$

   d. If 1.2 moles of CH_2Cl_2 and 2.1 moles of Cl_2 are mixed, what reagent will you run out of first? That is, what is the limiting reagent?

   If the 1.2 moles of CH_2Cl_2 react completely, they'll need 2.4 moles of Cl_2. Since you have less Cl_2 than that, you'll run out of Cl_2 first. Cl_2 is the limiting reagent.

2. a. If the instructions say to use 1.2 moles of CH_2Cl_2, how many grams of CH_2Cl_2 should you measure out?

First you calculate the molecular mass of CH_2Cl_2. The atomic masses are: C, 12; H, 1; and Cl, 35.5 (they'll be given for any problem of this kind). The molecular mass of CH_2Cl_2 is (12 x 1) + (1 x 2) + (35.5 x 2) = 85 grams/mole. To calculate the number of grams in 1.2 moles, you multiply. This way the units of moles cancel out to give you the answer in grams.

$$1.2\ \cancel{\text{moles}}\ \text{x}\ 85\ \text{grams}/\cancel{\text{mole}} = 102\ \text{grams}$$

b) You isolate 17.8 grams of CCl_4 from a reaction. How many moles of CCl_4 is this?

The molecular mass of CCl_4 is (12 x 1) + (35.5 x 4) = 154 g/mol. Then you divide:

$$(17.8\ \text{g})/(154\ \text{g/mol}) = 0.116\ \text{mol}$$

A slightly different equation form is shown in the first part of the next problem.

3. a. If one gram of CH_2Cl_2 reacts with an excess of Cl_2, what is the theoretical yield of HCl in moles?

In all yield problems, the central part of the problem has to do with moles. From the balanced equation, one mole of CH_2Cl_2 gives two moles of HCl. So first you have to convert the 1 g of CH_2Cl_2 to moles. Notice that the units cancel.

$$\frac{1\ \text{g}\ CH_2Cl_2}{\ } \quad \frac{(1\ \text{mole}\ CH_2Cl_2)}{(85\ \text{g}\ CH_2Cl_2)} = 0.012\ \text{mole}\ CH_2Cl_2$$

Then you convert the moles of CH_2Cl_2 to moles of HCl:

$$\frac{0.012\ \text{mole}\ CH_2Cl_2}{\ } \quad \frac{(2\ \text{moles HCl})}{(1\ \text{mole}\ CH_2Cl_2)} = 0.024\ \text{mole HCl}$$

b. If 10.2 grams of Cl_2 react with an excess of CH_2Cl_2, what is the theoretical yield of CCl_4 in moles?

Don't forget that the central feature is the moles conversion. You can combine the two steps in one, keeping in mind that the units have to cancel.

$$\frac{10.2\ \text{g}\ Cl_2}{\ } \quad \frac{(1\ \text{mole}\ Cl_2)}{(71\ \text{g}\ Cl_2)} \quad \frac{(1\ \text{mole}\ CCl_4)}{(2\ \text{moles}\ Cl_2)} = 0.072\ \text{mole}\ CCl_4$$

4. a. If 2.7 grams of CH_2Cl_2 reacts with an excess of (that is, more than enough) Cl_2 this way, what is the theoretical yield of CCl_4 in grams?

The central part of the problem lies in the moles conversion: 1 mole of CH_2Cl_2 gives 1 mole of CCl_4. Both ends of the problem have grams in them, so you have to convert grams to moles, then moles to moles, and finally moles to grams. That's the plan. Here's how you carry it out:

$$\frac{2.7\text{ g }CH_2Cl_2}{} \quad \frac{(1\text{ mole }CH_2Cl_2)}{(85\text{ g }CH_2Cl_2)} \quad \frac{(1\text{ mole }CCl_4)}{(1\text{ mole }CH_2Cl_2)} \quad \frac{(154\text{ g }CCl_4)}{(1\text{ mole }CCl_4)} = 4.89\text{ g }CCl_4$$

The theoretical yield of CCl_4 in grams is larger than the amount of CH_2Cl_2 used. The Cl_2 provides the rest of the mass. It's the number of moles that's important.

b. If you mix 8.9 g of CH_2Cl_2 with 11.1 g of Cl_2, what is the limiting reagent?

You do this in moles, of course. You have (8.9 g)/(85 g/mol) = 0.10 mole of CH_2Cl_2 and (11.1 g)/(71 g/mol) = 0.16 mole of Cl_2. But from the balanced equation, you need two moles of Cl_2 for each mole of CH_2Cl_2. That means you need 0.20 mole total of Cl_2 for the 0.10 mole of CH_2Cl_2 you have. Cl_2 is the limiting reagent.

c. If you mix 8.9 g of CH_2Cl_2 with 11.1 g of Cl_2, what is the theoretical yield of CCl_4 in grams?

From part (b) above, you know that Cl_2 is the limiting reagent, and you do your calculations from that. Don't forget: grams of starting material to moles, moles of starting material to moles of product, moles of product to grams of product. The equation is:

$$\frac{11.1\text{ g }Cl_2}{} \quad \frac{(1\text{ mole }Cl_2)}{(71\text{ g }Cl_2)} \quad \frac{(1\text{ mole }CCl_4)}{(2\text{ moles }Cl_2)} \quad \frac{(154\text{ g }CCl_4)}{(1\text{ mole }CCl_4)} = 12\text{ g }CCl_4$$

5. a. If you mix 8.9 g of CH_2Cl_2 with 27.1 g of Cl_2, what is the limiting reagent?

You do this in moles. You have (8.9 g)/(85 g/mol) = 0.10 mole of CH_2Cl_2, and (27.1 g)/(71 g/mol) = 0.38 mole of Cl_2. From the balanced equation, you need two moles of Cl_2 for each mole of CH_2Cl_2. That means you need 0.20 mole total of Cl_2 for the 0.10 mole of CH_2Cl_2 you have. You have more than enough Cl_2 this time, and CH_2Cl_2 is the limiting reagent.

b. If you mix 8.9 g of CH_2Cl_2 with 27.1 g of Cl_2, what is the theoretical yield of CCl_4 in grams?

From part (a) above, you know that CH_2Cl_2 is the limiting reagent, and you do your calculations from that. It's grams of starting material to moles, moles of starting material to moles of product, moles of product to grams of product. The equation is:

$$\frac{8.9\text{ g }CH_2Cl_2}{} \; \frac{(1\text{ mole }CH_2Cl_2)}{(85\text{ g }CH_2Cl_2)} \; \frac{(1\text{ mole }CCl_4)}{(1\text{ mole }CH_2Cl_2)} \; \frac{(154\text{ g }CCl_4)}{(1\text{ mole }CCl_4)} = 16\text{ g }CCl_4$$

c. If you mix 8.9 g of CH_2Cl_2 with 27.1 g of Cl_2, and get 13 g of CCl_4, what is the percent yield of CCl_4 in grams?

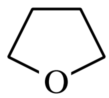
The percent yield is:

$$\frac{\text{Actual yield}}{\text{Theoretical Yield}} \times 100\% = \frac{13\text{ g}}{16\text{ g}} \times 100\% = 81\%$$

Note: In the sample questions on the previous pages, you are led to solve the final problem one step at a time. The problem is not always presented this way. Below you will find a sample test (B) that requires you to solve the steps without being prompted. You should be aware that this kind of problem statement also exists.

Name ____ Sample Yield Exercise A
Organic Chemistry 2220 Answer All Questions

Consider the following reaction:

 + 2 HCl ---> $ClCH_2CH_2CH_2CH_2Cl$ + H_2O

THF 1,4-dichlorobutane

You decide to run this reaction using 1.00 g of THF and 2.00 g of HCl. Atomic weights are: H, 1; C, 12; O, 16; Cl, 35.5. In answering the questions, show your work.

1. How many moles of THF are you using?

2. What is the limiting reagent?

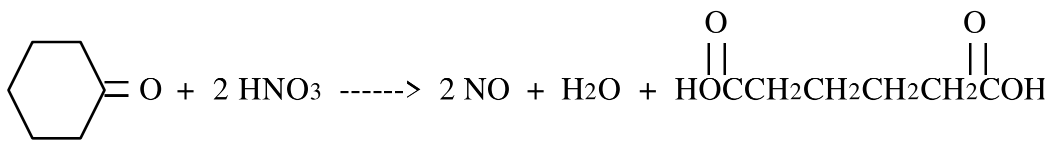
3. What is the theoretical yield of 1,4-dichlorobutane in moles?

4. What is the theoretical yield of 1,4-dichlorobutane in grams?

5. You actually run the reaction, and isolate 0.70 g of 1,4-dichlorobutane. What is your percent yield?

Name ____ Sample Yield Exercise B
Organic Chemistry 2220 Answer All Questions

Consider the following reaction:



1. When cyclohexanone is oxidized to adipic acid by nitric acid, NO is a byproduct. What mass of NO can be formed when 10.5 g of cyclohexanone is oxidized by an amount of nitric acid equivalent to 15.2 g of HNO_3? Show your work. Atomic weights are: H, 1.01; C, 12.0; N, 14.0; O, 16.0.

2. In a different run of the same reaction, the theoretical yield of NO was 9.49 g. The amount of NO actually obtained in that run was 5.25 g. What is the percent yield? Again, show your work.