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Percent Yield

Although we can write perfectly balanced equations to represent perfect reactions, the reactions themselves are often not perfect. A reaction does not always produce the quantity of products that the balanced equation seems to guarantee. This happens not because the equation is wrong but because reactions in the real world seldom produce perfect results.

As an example of an imperfect reaction, look again at the equation that shows the industrial production of ammonia.

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

In the manufacture of ammonia, it is nearly impossible to produce 2 mol (34.08 g) of NH_3 from the simple reaction of 1 mol (28.02 g) of N_2 and 3 mol (6.06 g) of H_2 because some ammonia molecules begin breaking down into N_2 and H_2 molecules as soon as they are formed.

There are several reasons that real-world reactions do not produce products at a yield of 100%. Some are simple mechanical reasons, such as:

- Reactants or products leak out, especially when they are gases.
- The reactants are not 100% pure.
- Some product is lost when it is purified.

There are also many chemical reasons, including:

- The products decompose back into reactants (as with the ammonia process).
- The products react to form different substances.
- Some of the reactants react in ways other than the one shown in the equation. These are called *side reactions.*
- The reaction occurs very slowly. This is especially true of reactions involving organic substances.

Chemists are very concerned with the yields of reactions because they must find ways to carry out reactions economically and on a large scale. If the yield of a reaction is too small, the products may not be competitive in the marketplace. If a reaction has only a 50% yield, it produces only 50% of the amount of product that it theoretically should. In this chapter, you will learn how to solve problems involving real-world reactions and percent yield.

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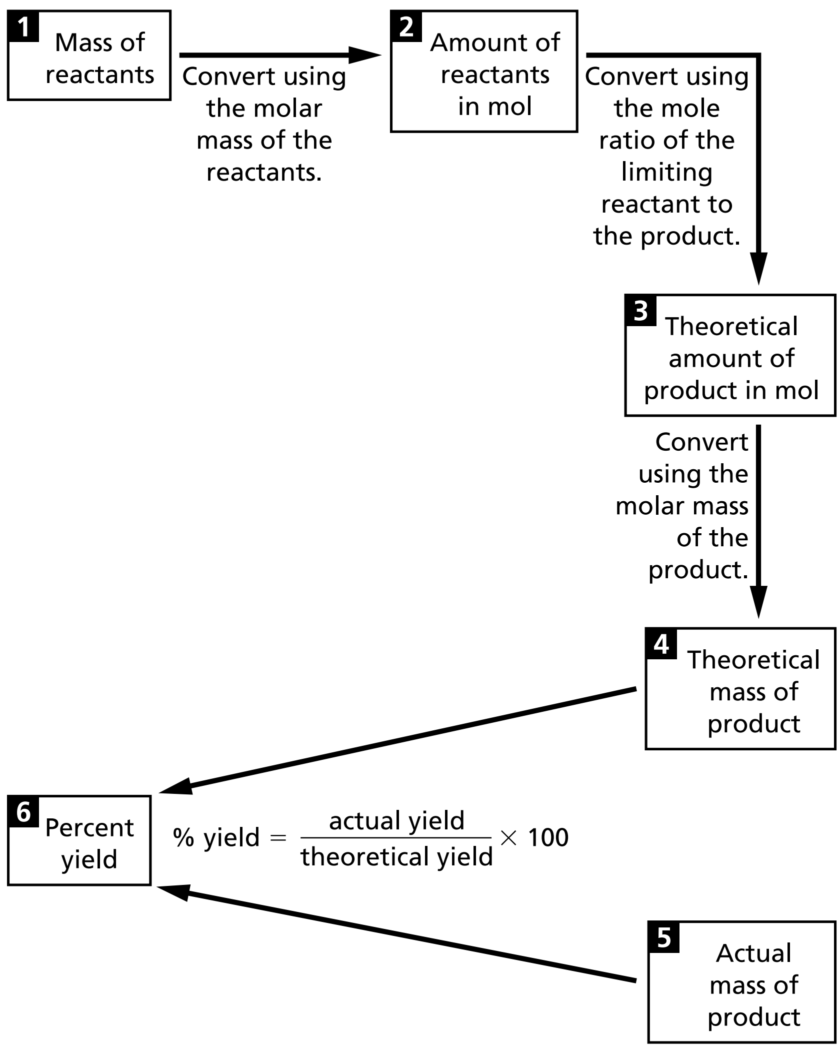
General Plan for Solving Percent-Yield Problems



SAMPLE PROBLEM 1

Dichlorine monoxide, Cl_2O is sometimes used as a powerful chlorinating agent in research. It can be produced by passing chlorine gas over heated mercury(II) oxide according to the following equation:

$$HgO + Cl_2 \rightarrow HgCl_2 + Cl_2O$$

What is the percent yield, if the quantity of reactants is sufficient to produce 0.86 g of Cl_2O but only 0.71 g is obtained?

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SOLUTION

1. *ANALYZE*

- *What is given in the problem?* — the balanced equation, the actual yield of Cl_2O, and the theoretical yield of Cl_2O
- *What are you asked to find?* — the percent yield of Cl_2O

Items	Data
Substance	Cl_2O
Mass available	NA*
Molar mass	NA
Amount of reactant	NA
Coefficient in balanced equation	NA
Actual yield	0.71 g
Theoretical yield (moles)	NA
Theoretical yield (grams)	0.86 g
Percent yield	?

* Although this table has many *Not Applicable* entries, you will need much of this information in other kinds of percent-yield problems.

2. *PLAN*

- *What steps are needed to calculate the percent yield of Cl_2O?* — Compute the ratio of the actual yield to the theoretical yield, and multiply by 100 to convert to a percentage.

4 Theoretical mass of Cl_2O in g → **6** Percent yield of Cl_2O

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

5 Actual mass of Cl_2O in g → **6** Percent yield of Cl_2O

$$\overbrace{\frac{\text{g } Cl_2O \text{ produced}}{\text{theoretical g } Cl_2O}}^{\frac{actual\ mass}{theoretical\ mass}} \times 100 = \text{percent yield}$$

3. *COMPUTE*

$$\frac{0.71 \text{ g } Cl_2O}{0.86 \text{ g } Cl_2O} \times 100 = 83\% \text{ yield}$$

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4. *EVALUATE*

- *Are the units correct?* — Yes; the ratio was converted to a percentage.
- *Is the number of significant figures correct?* — Yes; the number of significant figures is correct because the data were given to two significant figures.
- *Is the answer reasonable?* — Yes; 83% is about 5/6, which appears to be close to the ratio 0.71/0.86.

PRACTICE

1. Calculate the percent yield in each of the following cases:

a. theoretical yield is 50.0 g of product; actual yield is 41.9 g — *ans:* 83.8% yield

b. theoretical yield is 290 kg of product; actual yield is 270 kg — *ans:* 93% yield

c. theoretical yield is 6.05×10^4 kg of product; actual yield is 4.18×10^4 kg — *ans:* 69.1% yield

d. theoretical yield is 0.00192 g of product; actual yield is 0.00089 g — *ans:* 46% yield

SAMPLE PROBLEM 2

Acetylene, C_2H_2, can be used as an industrial starting material for the production of many organic compounds. Sometimes, it is first brominated to form 1,1,2,2-tetrabromoethane, $CHBr_2CHBr_2$, which can then be reacted in many different ways to make other substances. The equation for the bromination of acetylene follows:

$$\underset{\textit{acetylene}}{C_2H_2} + 2Br_2 \rightarrow \underset{\textit{1,1,2,2-tetrabromoethane}}{CHBr_2CHBr_2}$$

If 72.0 g of C_2H_2 reacts with excess bromine and 729 g of the product is recovered, what is the percent yield of the reaction?

SOLUTION

1. *ANALYZE*

- *What is given in the problem?* — the balanced equation, the mass of acetylene that reacts, and the mass of tetrabromoethane produced
- *What are you asked to find?* — the percent yield of tetrabromoethane

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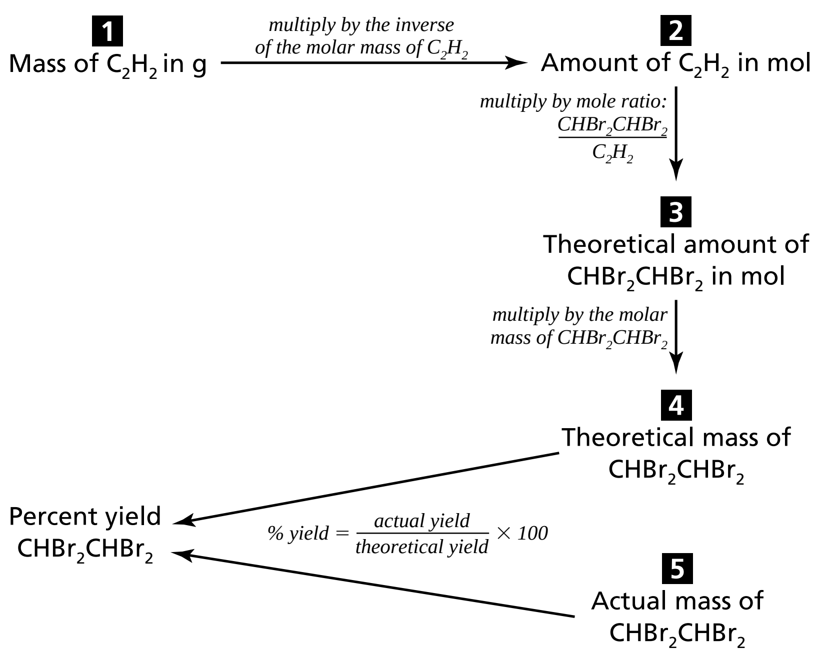
Items	Data	
Substance	C_2H_2	$CHBr_2CHBr_2$
Mass available	72.0 g available	NA
Molar mass*	26.04 g/mol	345.64 g/mol
Amount of reactant	?	NA
Coefficient in balanced equation	1	1
Actual yield	NA	729 g
Theoretical yield (moles)	NA	?
Theoretical yield (grams)	NA	?
Percent yield	NA	?

* determined from the periodic table

2. *PLAN*

- *What steps are needed to calculate the theoretical yield of tetrabromoethane?* Set up a stoichiometry calculation to find the amount of product that can be formed from the given amount of reactant.
- *What steps are needed to calculate the percent yield of tetrabromoethane.* Compute the ratio of the actual yield to the theoretical yield, and multiply by 100 to convert to a percentage.



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$$\overset{given}{g\ C_2H_2} \times \overset{\frac{1}{\text{molar mass } C_2H_2}}{\frac{1\text{ mol } C_2H_2}{26.04\text{ g } C_2H_2}} \times \overset{\text{mole ratio}}{\frac{1\text{ mol } CHBr_2CHBr_2}{1\text{ mol } C_2H_2}} \times \overset{\text{molar mass } CHBr_2CHBr_2}{\frac{345.64\text{ g } CHBr_2CHBr_2}{1\text{ mol } CHBr_2CHBr_2}}$$
$$= \text{theoretical g } CHBr_2CHBr_2$$

$$\overset{\frac{\text{actual grams}}{\text{theoretical grams}}}{\frac{\text{g } CHBr_2CHBr_2 \text{ produced}}{\text{theoretical g } CHBr_2CHBr_2}} \times 100 = \text{percent yield } CHBr_2CHBr_2$$

3. *COMPUTE*

$$72.0\ \cancel{g\ C_2H_2} \times \frac{1\ \cancel{mol\ C_2H_2}}{26.04\ \cancel{g\ C_2H_2}} \times \frac{1\ \cancel{mol\ CHBr_2CHBr_2}}{1\ \cancel{mol\ C_2H_2}}$$
$$\times \frac{345.64\text{ g } CHBr_2CHBr_2}{1\ \cancel{mol\ CHBr_2CHBr_2}} = 956\text{ g } CHBr_2CHBr_2$$

$$\frac{729\text{ g } CHBr_2CHBr_2}{956\text{ g } CHBr_2CHBr_2} \times 100 = 76.3\%\text{ yield}$$

4. *EVALUATE*

- *Are the units correct?* Yes; units canceled to give grams of $CHBr_2CHBr_2$. Also, the ratio was converted to a percentage.
- *Is the number of significant figures correct?* Yes; the number of significant figures is correct because the data were given to three significant figures.
- *Is the answer reasonable?* Yes; about 3 mol of acetylene were used and the theoretical yield is the mass of about 3 mol tetrabromoethane.

PRACTICE

1. In the commercial production of the element arsenic, arsenic(III) oxide is heated with carbon, which reduces the oxide to the metal according to the following equation:

$$2As_2O_3 + 3C \rightarrow 3CO_2 + 4As$$

a. If 8.87 g of As_2O_3 is used in the reaction and 5.33 g of As is produced, what is the percent yield? *ans:* 79.3% yield

b. If 67 g of carbon is used up in a different reaction and 425 g of As is produced, calculate the percent yield of this reaction. *ans:* 76% yield

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ADDITIONAL PROBLEMS

1. Ethyl acetate is a sweet-smelling solvent used in varnishes and fingernail-polish remover. It is produced industrially by heating acetic acid and ethanol together in the presence of sulfuric acid, which is added to speed up the reaction. The ethyl acetate is distilled off as it is formed. The equation for the process is as follows.

$$\underset{\textit{acetic acid}}{CH_3COOH} + \underset{\textit{ethanol}}{CH_3CH_2OH} \xrightarrow{H_2SO_4} \underset{\textit{ethyl acetate}}{CH_3COOCH_2CH_3} + H_2O$$

Determine the percent yield in the following cases:

a. 68.3 g of ethyl acetate should be produced but only 43.9 g is recovered.

b. 0.0419 mol of ethyl acetate is produced but 0.0722 mol is expected. (Hint: Percent yield can also be calculated by dividing the actual yield in moles by the theoretical yield in moles.)

c. 4.29 mol of ethanol is reacted with excess acetic acid, but only 2.98 mol of ethyl acetate is produced.

d. A mixture of 0.58 mol ethanol and 0.82 mol acetic acid is reacted and 0.46 mol ethyl acetate is produced. (Hint: What is the limiting reactant?)

2. Assume the following hypothetical reaction takes place.

$$2A + 7B \rightarrow 4C + 3D$$

Calculate the percent yield in each of the following cases:

a. The reaction of 0.0251 mol of A produces 0.0349 mol of C.

b. The reaction of 1.19 mol of A produces 1.41 mol of D.

c. The reaction of 189 mol of B produces 39 mol of D.

d. The reaction of 3500 mol of B produces 1700 mol of C.

3. Elemental phosphorus can be produced by heating calcium phosphate from rocks with silica sand (SiO_2) and carbon in the form of coke. The following reaction takes place.

$$Ca_3(PO_4)_2 + 3SiO_2 + 5C \rightarrow 3CaSiO_3 + 2P + 5CO$$

a. If 57 mol of $Ca_3(PO_4)_2$ is used and 101 mol of $CaSiO_3$ is obtained, what is the percent yield?

b. Determine the percent yield obtained if 1280 mol of carbon is consumed and 622 mol of $CaSiO_3$ is produced.

c. The engineer in charge of this process expects a yield of 81.5%. If 1.4×10^5 mol of $Ca_3(PO_4)_2$ is used, how many moles of phosphorus will be produced?

4. Tungsten (W) can be produced from its oxide by reacting the oxide with hydrogen at a high temperature according to the following equation:

$$WO_3 + 3H_2 \rightarrow W + 3H_2O$$

a. What is the percent yield if 56.9 g of WO_3 yields 41.4 g of tungsten?

b. How many moles of tungsten will be produced from 3.72 g of WO_3 if the yield is 92.0%?

c. A chemist carries out this reaction and obtains 11.4 g of tungsten. If the percent yield is 89.4%, what mass of WO_3 was used?

5. Carbon tetrachloride, CCl_4, is a solvent that was once used in large quantities in dry cleaning. Because it is a dense liquid that does not burn, it was also used in fire extinguishers. Unfortunately, its use was discontinued because it was found to be a carcinogen. It was manufactured by the following reaction:

$$CS_2 + 3Cl_2 \rightarrow CCl_4 + S_2Cl_2$$

The reaction was economical because the byproduct disulfur dichloride, S_2Cl_2, could be used by industry in the manufacture of rubber products and other materials.

a. What is the percent yield of CCl_4 if 719 kg is produced from the reaction of 410. kg of CS_2.

b. If 67.5 g of Cl_2 are used in the reaction and 39.5 g of S_2Cl_2 is produced, what is the percent yield?

c. If the percent yield of the industrial process is 83.3%, how many kilograms of CS_2 should be reacted to obtain 5.00×10^4 kg of CCl_4? How many kilograms of S_2Cl_2 will be produced, assuming the same yield for that product?

6. Nitrogen dioxide, NO_2, can be converted to dinitrogen pentoxide, N_2O_5, by reacting it with ozone, O_3. The reaction of NO_2 takes place according to the following equation:

$$2NO_2(g) + O_3(g) \rightarrow N_2O_5(s \text{ or } g) + O_2(g)$$

a. Calculate the percent yield for a reaction in which 0.38 g of NO_2 reacts and 0.36 g of N_2O_5 is recovered.

b. What mass of N_2O_5 will result from the reaction of 6.0 mol of NO_2 if there is a 61.1% yield in the reaction?

7. In the past, hydrogen chloride, HCl, was made using the *salt-cake* method as shown in the following equation:

$$2NaCl(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(s) + 2HCl(g)$$

If 30.0 g of NaCl and 0.250 mol of H_2SO_4 are available, and 14.6 g of HCl is made, what is the percent yield?

8. Cyanide compounds such as sodium cyanide, NaCN, are especially useful in gold refining because they will react with gold to form a stable compound that can then be separated and broken down to retrieve the gold. Ore containing only small quantities of gold can be used in this form of "chemical mining." The equation for the reaction follows.

$$4Au + 8NaCN + 2H_2O + O_2 \rightarrow 4NaAu(CN)_2 + 4NaOH$$

a. What percent yield is obtained if 410 g of gold produces 540 g of $NaAu(CN)_2$?

b. Assuming a 79.6% yield in the conversion of gold to $NaAu(CN)_2$, what mass of gold would produce 1.00 kg of $NaAu(CN)_2$?

c. Given the conditions in (b), what mass of gold ore that is 0.001% gold would be needed to produce 1.00 kg of $NaAu(CN)_2$?

9. Diiodine pentoxide is useful in devices such as respirators because it reacts with the dangerous gas carbon monoxide, CO, to produce relatively harmless CO_2 according to the following equation:

$$I_2O_5 + 5CO \rightarrow I_2 + 5CO_2$$

a. In testing a respirator, 2.00 g of carbon monoxide gas is passed through diiodine pentoxide. Upon analyzing the results, it is found that 3.17 g of I_2 was produced. Calculate the percent yield of the reaction.

b. Assuming that the yield in (a) resulted because some of the CO did not react, calculate the mass of CO that passed through.

10. Sodium hypochlorite, NaClO, the main ingredient in household bleach, is produced by bubbling chlorine gas through a strong lye (sodium hydroxide, NaOH) solution. The following equation shows the reaction that occurs.

$$2NaOH(aq) + Cl_2(g) \rightarrow NaCl(aq) + NaClO(aq) + H_2O(l)$$

a. What is the percent yield of the reaction if 1.2 kg of Cl_2 reacts to form 0.90 kg of NaClO?

b. If a plant operator wants to make 25 metric tons of NaClO per day at a yield of 91.8%, how many metric tons of chlorine gas must be on hand each day?

c. What mass of NaCl is formed per mole of chlorine gas at a yield of 81.8%?

d. At what rate in kg per hour must NaOH be replenished if the reaction produces 370 kg/h of NaClO at a yield of 79.5%? Assume that all of the NaOH reacts to produce this yield.

11. Magnesium burns in oxygen to form magnesium oxide. However, when magnesium burns in air, which is only about 1/5 oxygen, side reactions form other products, such as magnesium nitride, Mg_3N_2.

a. Write a balanced equation for the burning of magnesium in oxygen.

b. If enough magnesium burns in air to produce 2.04 g of magnesium oxide but only 1.79 g is obtained, what is the percent yield?

c. Magnesium will react with pure nitrogen to form the nitride, Mg_3N_2. Write a balanced equation for this reaction.

d. If 0.097 mol of Mg react with nitrogen and 0.027 mol of Mg_3N_2 is produced, what is the percent yield of the reaction?

12. Some alcohols can be converted to organic acids by using sodium dichromate and sulfuric acid. The following equation shows the reaction of 1-propanol to propanoic acid.

$$3CH_3CH_2CH_2OH + 2Na_2Cr_2O_7 + 8H_2SO_4 \rightarrow 3CH_3CH_2COOH + 2Cr_2(SO_4)_3 + 2Na_2SO_4 + 11H_2O$$

a. If 0.89 g of 1-propanol reacts and 0.88 g of propanoic acid is produced, what is the percent yield?

b. A chemist uses this reaction to obtain 1.50 mol of propanoic acid. The reaction consumes 136 g of propanol. Calculate the percent yield.

c. Some 1-propanol of uncertain purity is used in the reaction. If 116 g of $Na_2Cr_2O_7$ are consumed in the reaction and 28.1 g of propanoic acid are produced, what is the percent yield?

13. Acrylonitrile, $C_3H_3N(g)$, is an important ingredient in the production of various fibers and plastics. Acrylonitrile is produced from the following reaction:

$$C_3H_6(g) + NH_3(g) + O_2(g) \rightarrow C_3H_3N(g) + H_2O(g)$$

If 850. g of C_3H_6 is mixed with 300. g of NH_3 and unlimited O_2, to produce 850. g of acrylonitrile, what is the percent yield? You must first balance the equation

14. Methanol, CH_3OH, is frequently used in race cars as fuel. It is produced as the sole product of the combination of carbon monoxide gas and hydrogen gas.

a. If 430. kg of hydrogen react, what mass of methanol could be produced?

b. If 3.12×10^3 kg of methanol are actually produced, what is the percent yield?

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15. The compound, $C_6H_{16}N_2$, is one of the starting materials in the production of nylon. It can be prepared from the following reaction involving adipic acid, $C_6H_{10}O_4$:

$$C_6H_{10}O_4(l) + 2NH_3(g) + 4H_2(g) \rightarrow C_6H_{16}N_2(l) + 4H_2O$$

What is the percent yield if 750. g of adipic acid results in the production of 578 g of $C_6H_{16}N_2$?

16. Plants convert carbon dioxide to oxygen during photosynthesis according to the following equation:

$$CO_2 + H_2O \rightarrow C_6H_{12}O_6 + O_2$$

Balance this equation, and calculate how much oxygen would be produced if 1.37×10^4 g of carbon dioxide reacts with a percent yield of 63.4%.

17. Lime, CaO, is frequently added to streams and lakes which have been polluted by acid rain. The calcium oxide reacts with the water to form a base that can neutralize the acid as shown in the following reaction:

$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$$

If 2.67×10^2 mol of base are needed to neutralize the acid in a lake, and the above reaction has a percent yield of 54.3%, what is the mass, in kilograms, of lime that must be added to the lake?