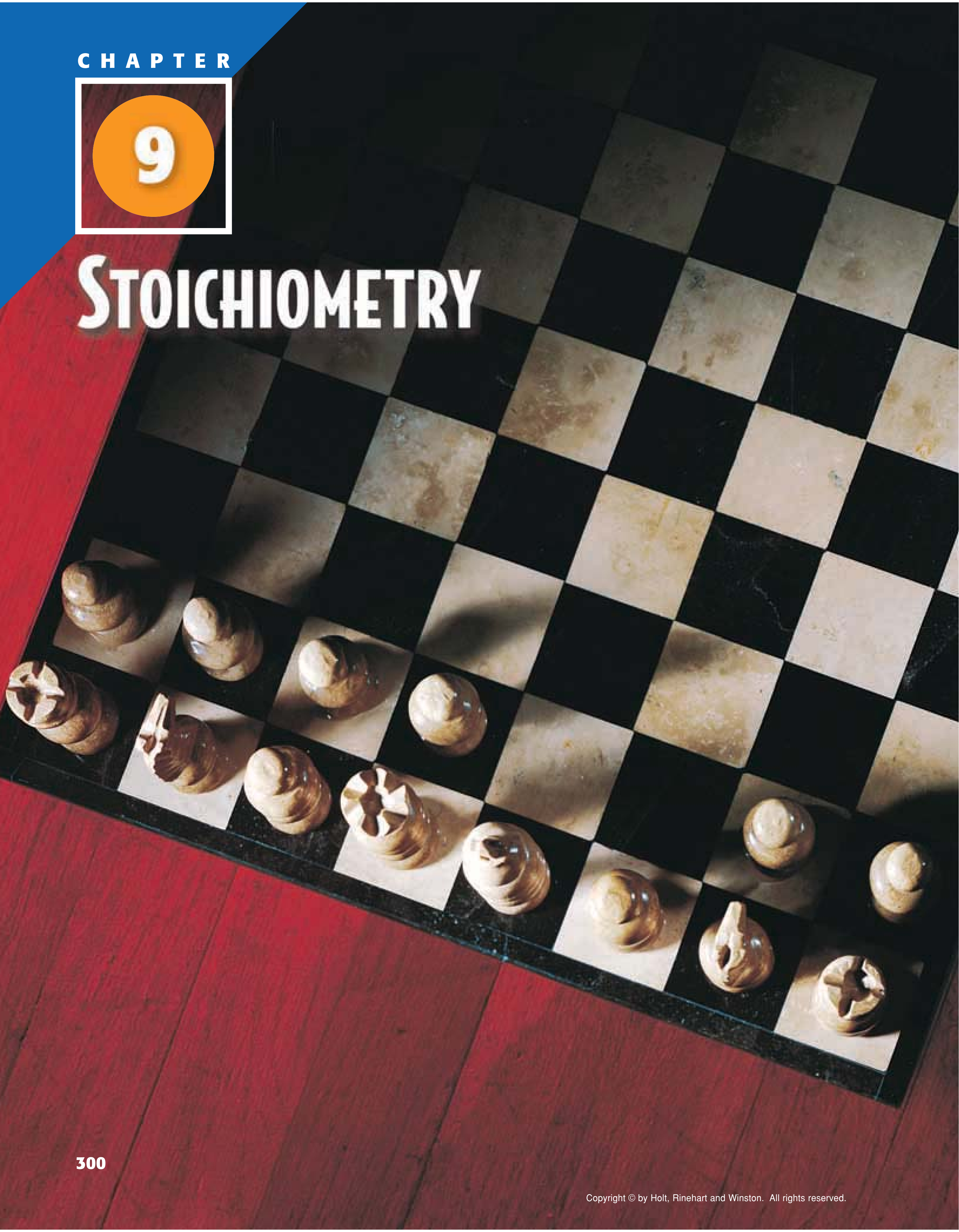
CHAPTER

9

STOICHIOMETRY



To play a standard game of chess, each side needs the proper number of pieces and pawns. Unless you find all of them—a king, a queen, two bishops, two knights, two rooks, and eight pawns—you cannot start the game. In chemical reactions, if you do not have every reactant, you will not be able to start the reaction. In this chapter you will look at amounts of reactants present and calculate the amounts of other reactants or products that are involved in the reaction.

START-UP ACTIVITY

All Used Up

PROCEDURE

1. Use a **balance** to find the mass of **8 nuts** and the mass of **5 bolts**.
2. Attach 1 nut (N) to 1 bolt (B) to assemble a nut-bolt (NB) model. Make as many NB models as you can. Record the number of models formed, and record which material was used up. Take the models apart.
3. Attach 2 nuts to 1 bolt to assemble a nut-nut-bolt (N_2B) model. Make as many N_2B models as you can. Record the number of models formed, and record which material was used up. Take the models apart.

ANALYSIS

1. Using the masses of the starting materials (the nuts and the bolts), could you predict which material would be used up first? Explain.
2. Write a balanced equation for the “reaction” that forms NB. How can this equation help you predict which component runs out?
3. Write a balanced equation for the “reaction” that forms N_2B. How can this equation help you predict which component runs out?
4. If you have 18 bolts and 26 nuts, how many models of NB could you make? of N_2B?

Pre-Reading Questions

1. **A recipe calls for one cup of milk and three eggs per serving. You quadruple the recipe because you're expecting guests. How much milk and eggs do you need?**
2. **A bicycle mechanic has 10 frames and 16 wheels in the shop. How many complete bicycles can he assemble using these parts?**
3. **List at least two conversion factors that relate to the mole.**

CONTENTS 9

SECTION

1 Calculating Quantities in Reactions

KEY TERMS

- **stoichiometry**

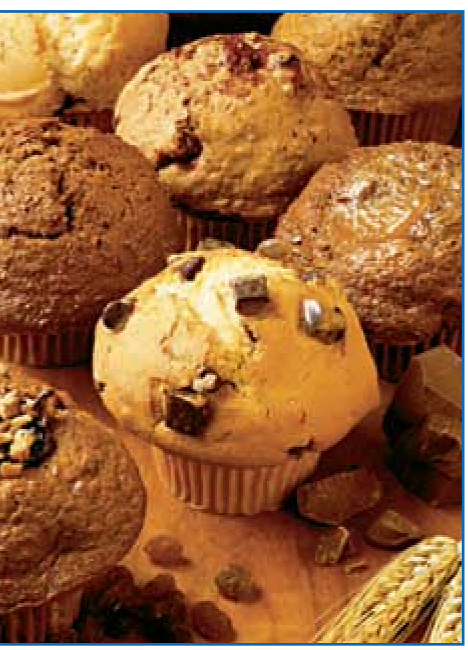
OBJECTIVES

1. **Use** proportional reasoning to determine mole ratios from a balanced chemical equation.
2. **Explain** why mole ratios are central to solving stoichiometry problems.
3. **Solve** stoichiometry problems involving mass by using molar mass.
4. **Solve** stoichiometry problems involving the volume of a substance by using density.
5. **Solve** stoichiometry problems involving the number of particles of a substance by using Avogadro's number.

Balanced Equations Show Proportions

If you wanted homemade muffins, like the ones in **Figure 1,** you could make them yourself—if you had the right things. A recipe for muffins shows how much of each ingredient you need to make 12 muffins. It also shows the proportions of those ingredients. If you had just a little flour on hand, you could determine how much of the other things you should use to make great muffins. The proportions also let you adjust the amounts to make enough muffins for all your classmates.

Figure 1
In using a recipe to make muffins, you are using proportions to determine how much of each ingredient is needed.



A balanced chemical equation is very similar to a recipe in that the coefficients in the balanced equation show the proportions of the reactants and products involved in the reaction. For example, consider the reaction for the synthesis of water.

$$2H_2 + O_2 \rightarrow 2H_2O$$

On a very small scale, the coefficients in a balanced equation represent the numbers of particles for each substance in the reaction. For the equation above, the coefficients show that two molecules of hydrogen react with one molecule of oxygen and form two molecules of water.

Calculations that involve chemical reactions use the proportions from balanced chemical equations to find the quantity of each reactant and product involved. As you learn how to do these calculations in this section, you will assume that each reaction goes to completion. In other words, all of the given reactant changes into product. For each problem in this section, assume that there is more than enough of all other reactants to completely react with the reactant given. Also assume that every reaction happens perfectly, so that no product is lost during collection. As you will learn in the next section, this usually is not the case.

Relative Amounts in Equations Can Be Expressed in Moles

Just as you can interpret equations in terms of particles, you can interpret them in terms of moles. The coefficients in a balanced equation also represent the moles of each substance. For example, the equation for the synthesis of water shows that 2 mol H_2 react with 1 mol O_2 to form 2 mol H_2O. Look at the equation below.

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$$

This equation shows that 2 molecules C_8H_{18} react with 25 molecules O_2 to form 16 molecules CO_2 and 18 molecules H_2O. And because Avogadro's number links molecules to moles, the equation also shows that 2 mol C_8H_{18} react with 25 mol O_2 to form 16 mol CO_2 and 18 mol H_2O.

In this chapter you will learn to determine how much of a reactant is needed to produce a given quantity of product, or how much of a product is formed from a given quantity of reactant. The branch of chemistry that deals with quantities of substances in chemical reactions is known as **stoichiometry.**

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stoichiometry

the proportional relationship between two or more substances during a chemical reaction

The Mole Ratio Is the Key

If you normally buy a lunch at school each day, how many times would you need to "brown bag" it if you wanted to save enough money to buy a CD player? To determine the answer, you would use the units of dollars to bridge the gap between a CD player and school lunches. In stoichiometry problems involving equations, the unit that bridges the gap between one substance and another is the mole.

The coefficients in a balanced chemical equation show the relative numbers of moles of the substances in the reaction. As a result, you can use the coefficients in conversion factors called *mole ratios*. Mole ratios bridge the gap and can convert from moles of one substance to moles of another, as shown in **Skills Toolkit 1.**

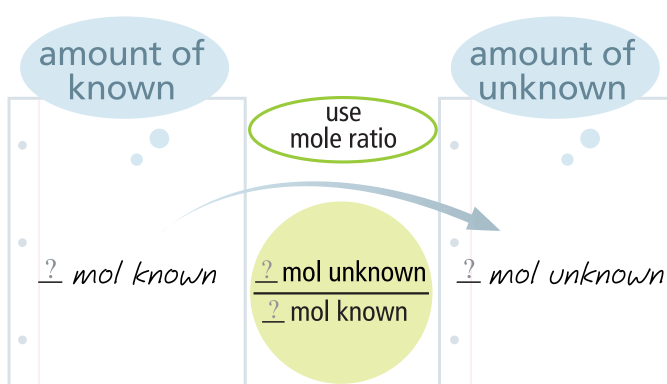
SKILLS Toolkit 1

Converting Between Amounts in Moles

1. Identify the amount in moles that you know from the problem.

2. Using coefficients from the balanced equation, set up the mole ratio with the known substance on bottom and the unknown substance on top.

3. Multiply the original amount by the mole ratio.



SAMPLE PROBLEM A

Using Mole Ratios

Consider the reaction for the commercial preparation of ammonia.

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

How many moles of hydrogen are needed to prepare 312 moles of ammonia?

1 Gather information.

- amount of NH_3 = 312 mol
- amount of H_2 = ? mol
- From the equation: 3 mol H_2 = 2 mol NH_3.

2 Plan your work.

The mole ratio must cancel out the units of mol NH_3 given in the problem and leave the units of mol H_2. Therefore, the mole ratio is

$$\frac{3 \text{ mol } H_2}{2 \text{ mol } NH_3}$$

3 Calculate.

$$? \text{ mol } H_2 = 312 \cancel{\text{mol } NH_3} \times \frac{3 \text{ mol } H_2}{2 \cancel{\text{mol } NH_3}} =$$

$$468 \text{ mol } H_2$$

4 Verify your result.

- The answer is larger than the initial number of moles of ammonia. This is expected, because the conversion factor is greater than one.
- The number of significant figures is correct because the coefficients 3 and 2 are considered to be exact numbers.

PRACTICE

PROBLEM SOLVING SKILL

1 Calculate the amounts requested if 1.34 mol H_2O_2 completely react according to the following equation.

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

a. moles of oxygen formed
b. moles of water formed

2 Calculate the amounts requested if 3.30 mol Fe_2O_3 completely react according to the following equation.

$$Fe_2O_3 + 2Al \rightarrow 2Fe + Al_2O_3$$

a. moles of aluminum needed
b. moles of iron formed
c. moles of aluminum oxide formed

PRACTICE HINT

The mole ratio must always have the unknown substance on top and the substance given in the problem on bottom for units to cancel correctly.

Getting into Moles and Getting out of Moles

Substances are usually measured by mass or volume. As a result, before using the mole ratio you will often need to convert between the units for mass and volume and the unit *mol*. Yet each stoichiometry problem has the step in which moles of one substance are converted into moles of a second substance using the mole ratio from the balanced chemical equation. Follow the steps in **Skills Toolkit 2** to understand the process of solving stoichiometry problems.

The thought process in solving stoichiometry problems can be broken down into three basic steps. First, change the units you are given into moles. Second, use the mole ratio to determine moles of the desired substance. Third, change out of moles to whatever unit you need for your final answer. And if you are given moles in the problem or need moles as an answer, just skip the first step or the last step! As you continue reading, you will be reminded of the conversion factors that involve moles.

Solving Stoichiometry Problems

You can solve all types of stoichiometry problems by following the steps outlined below.

1. Gather information.

- If an equation is given, make sure the equation is balanced. If no equation is given, write a balanced equation for the reaction described.
- Write the information provided for the given substance. If you are not given an amount in moles, determine the information you need to change the given units into moles and write it down.
- Write the units you are asked to find for the unknown substance. If you are not asked to find an amount in moles, determine the information you need to change moles into the desired units, and write it down.
- Write an equality using substances and their coefficients that shows the relative amounts of the substances from the balanced equation.

2. Plan your work.

- Think through the three basic steps used to solve stoichiometry problems: change to moles, use the mole ratio, and change out of moles. Know which conversion factors you will use in each step.
- Write the mole ratio you will use in the form:

$$\frac{\text{moles of unknown substance}}{\text{moles of given substance}}$$

3. Calculate.

- Write a question mark with the units of the answer followed by an equals sign and the quantity of the given substance.
- Write the conversion factors—including the mole ratio—in order so that you change the units of the given substance to the units needed for the answer.
- Cancel units and check that the remaining units are the required units of the unknown substance.
- When you have finished your calculations, round off the answer to the correct number of significant figures. In the examples in this book, only the final answer is rounded off.
- Report your answer with correct units and with the name or formula of the substance.

4. Verify your result.

- Verify your answer by estimating. You could round off the numbers in the setup in step 3 and make a quick calculation. Or you could compare conversion factors in the setup and decide whether the answer should be bigger or smaller than the initial value.
- Make sure your answer is reasonable. For example, imagine that you calculate that 725 g of a reactant is needed to form 5.3 mg (0.0053 g) of a product. The large difference in these quantities should alert you that there may be an error and that you should double-check your work.

Figure 2
These tanks store ammonia for use as fertilizer. Stoichiometry is used to determine the amount of ammonia that can be made from given amounts of H_2 and N_2.

Problems Involving Mass, Volume, or Particles

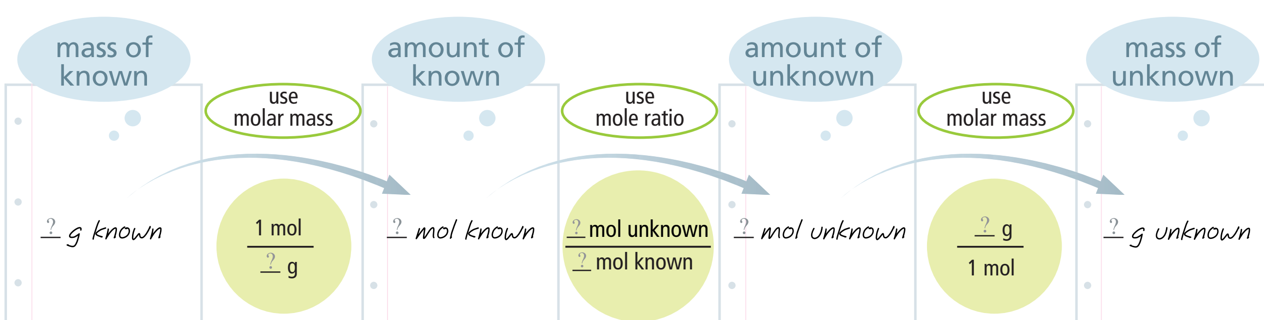
Figure 2 shows a few of the tanks used to store the millons of metric tons of ammonia made each year in the United States. Stoichiometric calculations are used to determine how much of the reactants are needed and how much product is expected. However, the calculations do not start and end with moles. Instead, other units, such as liters or grams, are used. Mass, volume, or number of particles can all be used as the starting and ending quantities of stoichiometry problems. Of course, the key to each of these problems is the mole ratio.

For Mass Calculations, Use Molar Mass

The conversion factor for converting between mass and amount in moles is the molar mass of the substance. The molar mass is the sum of atomic masses from the periodic table for the atoms in a substance. **Skills Toolkit 3** shows how to use the molar mass of each substance involved in a stoichiometry problem. Notice that the problem is a three-step process. The mass in grams of the given substance is converted into moles. Next, the mole ratio is used to convert into moles of the desired substance. Finally, this amount in moles is converted into grams.

3 SKILLS Toolkit

Solving Mass-Mass Problems



SAMPLE PROBLEM B

Problems Involving Mass

What mass of NH_3 can be made from 1221 g H_2 and excess N_2?

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

1 Gather information.

- mass of H_2 = 1221 g H_2
- molar mass of H_2 = 2.02 g/mol
- mass of NH_3 = ? g NH_3
- molar mass of NH_3 = 17.04 g/mol
- From the balanced equation: 3 mol H_2 = 2 mol NH_3.

2 Plan your work.

- To change grams of H_2 to moles, use the molar mass of H_2.
- The mole ratio must cancel out the units of mol H_2 given in the problem and leave the units of mol NH_3. Therefore, the mole ratio is

$$\frac{2\text{ mol }NH_3}{3\text{ mol }H_2}$$

- To change moles of NH_3 to grams, use the molar mass of NH_3.

3 Calculate.

$$?\text{ g }NH_3 = 1221\ \cancel{\text{g }H_2} \times \frac{1\ \cancel{\text{mol }H_2}}{2.02\ \cancel{\text{g }H_2}} \times \frac{2\ \cancel{\text{mol }NH_3}}{3\ \cancel{\text{mol }H_2}} \times \frac{17.04\text{ g }NH_3}{1\ \cancel{\text{mol }NH_3}} =$$

$$6867\text{ g }NH_3$$

4 Verify your result.

The units cancel to give the correct units for the answer. Estimating shows the answer should be about 6 times the original mass.

PRACTICE HINT

Remember to check both the units and the substance when canceling. For example, 1221 g H_2 cannot be converted to moles by multiplying by 1 mol NH_3/17.04 g NH_3. The units of grams in each one cannot cancel because they involve different substances.

PRACTICE

Use the equation below to answer the questions that follow.

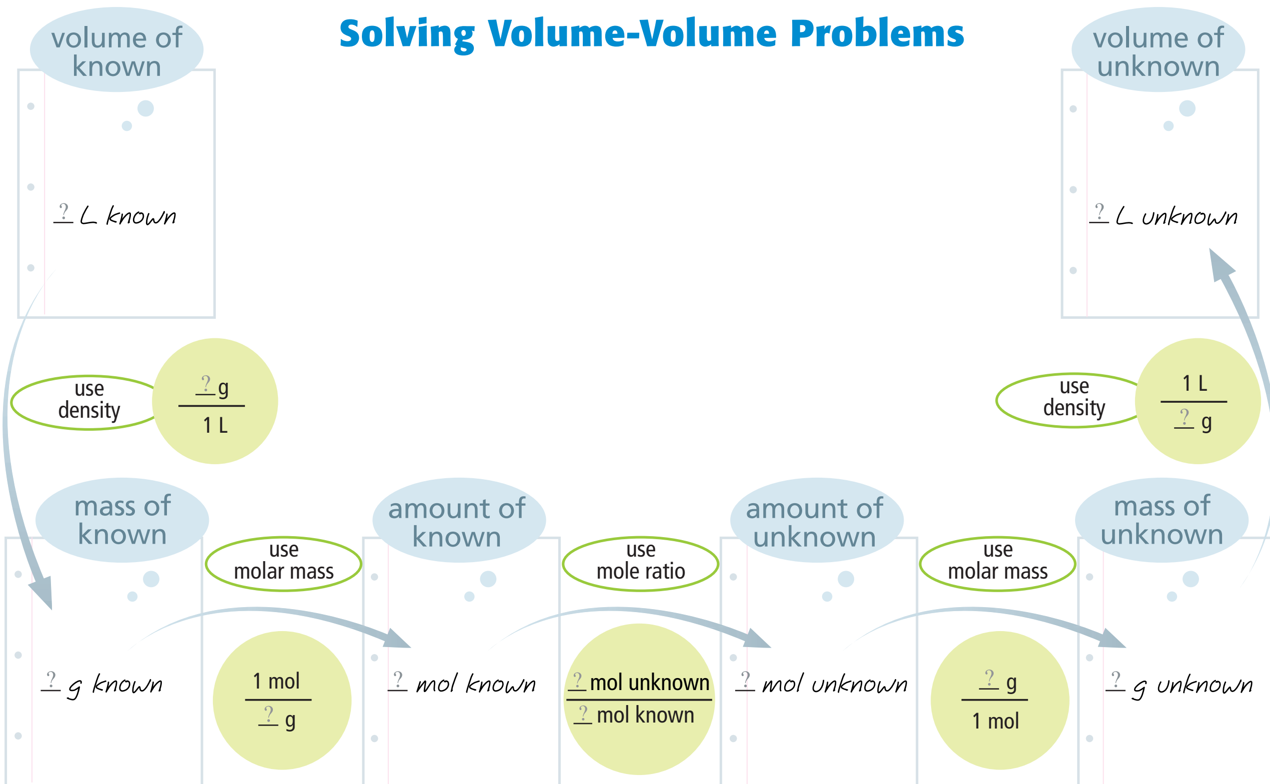
$$Fe_2O_3 + 2Al \longrightarrow 2Fe + Al_2O_3$$

1. How many grams of Al are needed to completely react with 135 g Fe_2O_3?
2. How many grams of Al_2O_3 can form when 23.6 g Al react with excess Fe_2O_3?
3. How many grams of Fe_2O_3 react with excess Al to make 475 g Fe?
4. How many grams of Fe will form when 97.6 g Al_2O_3 form?

PROBLEM SOLVING SKILL

4 SKILLS Toolkit

Solving Volume-Volume Problems



For Volume, You Might Use Density and Molar Mass

When reactants are liquids, they are almost always measured by volume. So, to do calculations involving liquids, you add two more steps to the sequence of mass-mass problems—the conversions of volume to mass and of mass to volume. Five conversion factors—two densities, two molar masses, and a mole ratio—are needed for this type of calculation, as shown in **Skills Toolkit 4.**

To convert from volume to mass or from mass to volume of a substance, use the density of the substance as the conversion factor. Keep in mind that the units you want to cancel should be on the bottom of your conversion factor.

There are ways other than density to include volume in stoichiometry problems. For example, if a substance in the problem is a gas at standard temperature and pressure (STP), use the molar volume of a gas to change directly between volume of the gas and moles. The molar volume of a gas is 22.41 L/mol for any gas at STP. Also, if a substance in the problem is in aqueous solution, then use the concentration of the solution to convert the volume of the solution to the moles of the substance dissolved. This procedure is especially useful when you perform calculations involving the reaction between an acid and a base. Of course, even in these problems, the basic process remains the same: change to moles, use the mole ratio, and change to the desired units.

SAMPLE PROBLEM C

Problems Involving Volume

What volume of H_3PO_4 forms when 56 mL $POCl_3$ completely react? (density of $POCl_3$ = 1.67 g/mL; density of H_3PO_4 = 1.83 g/mL)

$$POCl_3(l) + 3H_2O(l) \longrightarrow H_3PO_4(l) + 3HCl(g)$$

1 Gather information.

- volume $POCl_3$ = 56 mL $POCl_3$
- density $POCl_3$ = 1.67 g/mL
- molar mass $POCl_3$ = 153.32 g/mol
- volume H_3PO_4 = ?
- density H_3PO_4 = 1.83 g/mL
- molar mass H_3PO_4 = 98.00 g/mol
- From the equation: 1 mol $POCl_3$ = 1 mol H_3PO_4.

2 Plan your work.

- To change milliliters of $POCl_3$ to moles, use the density of $POCl_3$ followed by its molar mass.
- The mole ratio must cancel out the units of mol $POCl_3$ given in the problem and leave the units of mol H_3PO_4. Therefore, the mole ratio is

$$\frac{1 \text{ mol } H_3PO_4}{1 \text{ mol } POCl_3}$$

- To change out of moles of H_3PO_4 into milliliters, use the molar mass of H_3PO_4 followed by its density.

3 Calculate.

$$? \text{ mL } H_3PO_4 = 56 \cancel{\text{ mL } POCl_3} \times \frac{1.67 \cancel{\text{ g } POCl_3}}{1 \cancel{\text{ mL } POCl_3}} \times \frac{1 \cancel{\text{ mol } POCl_3}}{153.32 \cancel{\text{ g } POCl_3}} \times$$

$$\frac{1 \cancel{\text{ mol } H_3PO_4}}{1 \cancel{\text{ mol } POCl_3}} \times \frac{98.00 \cancel{\text{ g } H_3PO_4}}{1 \cancel{\text{ mol } H_3PO_4}} \times \frac{1 \text{ mL } H_3PO_4}{1.83 \cancel{\text{ g } H_3PO_4}} = 33 \text{ mL } H_3PO_4$$

4 Verify your result.

The units of the answer are correct. Estimating shows the answer should be about two-thirds of the original volume.

PRACTICE HINT

Do not try to memorize the exact steps of every type of problem. For long problems like these, you might find it easier to break the problem into three steps rather than solving all at once. Remember that whatever you are given, you need to change to moles, then use the mole ratio, then change out of moles to the desired units.

PRACTICE

PROBLEM SOLVING SKILL

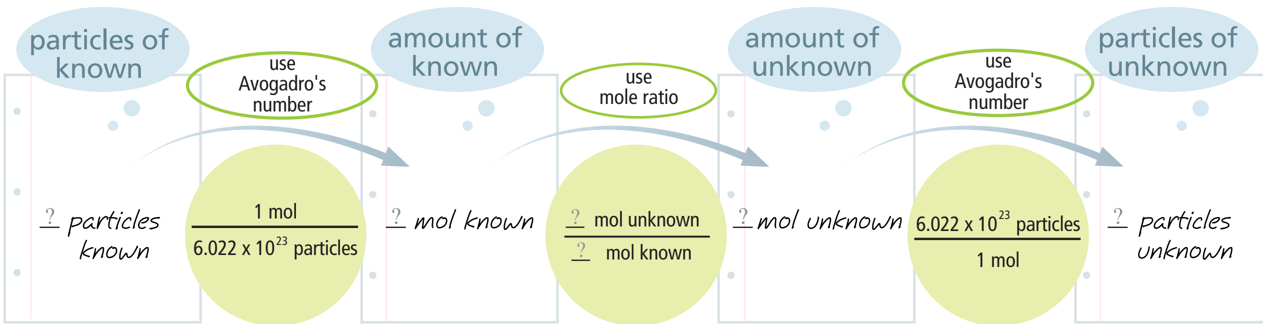
Use the densities and balanced equation provided to answer the questions that follow. (density of C_5H_{12} = 0.620 g/mL; density of C_5H_8 = 0.681 g/mL; density of H_2 = 0.0899 g/L)

$$C_5H_{12}(l) \longrightarrow C_5H_8(l) + 2H_2(g)$$

1. How many milliliters of C_5H_8 can be made from 366 mL C_5H_{12}?
2. How many liters of H_2 can form when 4.53×10^3 mL C_5H_8 form?
3. How many milliliters of C_5H_{12} are needed to make 97.3 mL C_5H_8?
4. How many milliliters of H_2 can be made from 1.98×10^3 mL C_5H_{12}?

5 SKILLS Toolkit

Solving Particle Problems



Topic Link

Refer to the chapter "The Mole and Chemical Composition" for more information about Avogadro's number and molar mass.

For Number of Particles, Use Avogadro's Number

Skills Toolkit 5 shows how to use Avogadro's number, 6.022×10^{23} particles/mol, in stoichiometry problems. If you are given particles and asked to find particles, Avogadro's number cancels out! For this calculation you use only the coefficients from the balanced equation. In effect, you are interpreting the equation in terms of the number of particles again.

PRACTICE HINT

Expect more problems like this one that do not exactly follow any single **Skills Toolkit** in this chapter. These problems will combine steps from one or more problems, but all will still use the mole ratio as the key step.

SAMPLE PROBLEM D

Problems Involving Particles

How many grams of C_5H_8 form from 1.89×10^{24} molecules C_5H_{12}?

$$C_5H_{12}(l) \longrightarrow C_5H_8(l) + 2H_2(g)$$

1 Gather information.

- quantity of $C_5H_{12} = 1.89 \times 10^{24}$ molecules
- Avogadro's number $= 6.022 \times 10^{23}$ molecules/mol
- mass of $C_5H_8 = ?$ g C_5H_8
- molar mass of $C_5H_8 = 68.13$ g/mol
- From the balanced equation: 1 mol C_5H_{12} = 1 mol C_5H_8.

2 Plan your work.

Set up the problem using Avogadro's number to change to moles, then use the mole ratio, and finally use the molar mass of C_5H_8 to change to grams.

3 Calculate.

$$? \text{ g } C_5H_8 = 1.89 \times 10^{24} \cancel{\text{molecules } C_5H_{12}} \times \frac{1 \cancel{\text{mol } C_5H_{12}}}{6.022 \times 10^{23} \cancel{\text{molecules } C_5H_{12}}} \times$$

$$\frac{1 \cancel{\text{mol } C_5H_8}}{1 \cancel{\text{mol } C_5H_{12}}} \times \frac{68.13 \text{ g } C_5H_8}{1 \cancel{\text{mol } C_5H_8}} = 214 \text{ g } C_5H_8$$

4 Verify your result.

The units cancel correctly, and estimating gives 210.

PRACTICE

PROBLEM SOLVING SKILL

Use the equation provided to answer the questions that follow.

$$Br_2(l) + 5F_2(g) \longrightarrow 2BrF_5(l)$$

1 How many molecules of BrF_5 form when 384 g Br_2 react with excess F_2?

2 How many molecules of Br_2 react with 1.11×10^{20} molecules F_2?

Many Problems, Just One Solution

Although you could be given many different problems, the solution boils down to just three steps. Take whatever you are given, and find a way to change it into moles. Then, use a mole ratio from the balanced equation to get moles of the second substance. Finally, find a way to convert the moles into the units that you need for your final answer.

1 Section Review

UNDERSTANDING KEY IDEAS

1. What conversion factor is present in almost all stoichiometry calculations?
2. For a given substance, what information links mass to moles? number of particles to moles?
3. What conversion factor will change moles CO_2 to grams CO_2? moles H_2O to molecules H_2O?

PRACTICE PROBLEMS

4. Use the equation below to answer the questions that follow.

   $$Br_2 + Cl_2 \longrightarrow 2BrCl$$

   a. How many moles of BrCl form when 2.74 mol Cl_2 react with excess Br_2?

   b. How many grams of BrCl form when 239.7 g Cl_2 react with excess Br_2?

   c. How many grams of Br_2 are needed to react with 4.53×10^{25} molecules Cl_2?

5. The equation for burning C_2H_2 is

   $$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(g)$$

   a. If 15.9 L C_2H_2 react at STP, how many moles of CO_2 are produced? (Hint: At STP, 1 mol = 22.41 L for any gas.)

   b. How many milliliters of CO_2 (density = 1.977 g/L) can be made when 59.3 mL O_2 (density = 1.429 g/L) react?

CRITICAL THINKING

6. Why do you need to use amount in moles to solve stoichiometry problems? Why can't you just convert from mass to mass?
7. LiOH and NaOH can each react with CO_2 to form the metal carbonate and H_2O. These reactions can be used to remove CO_2 from the air in a spacecraft.

   a. Write a balanced equation for each reaction.

   b. Calculate the grams of NaOH and of LiOH that remove 288 g CO_2 from the air.

   c. NaOH is less expensive per mole than LiOH. Based on your calculations, explain why LiOH is used during shuttle missions rather than NaOH.

SECTION

2 Limiting Reactants and Percentage Yield

KEY TERMS

- **limiting reactant**
- **excess reactant**
- **actual yield**

OBJECTIVES

1. **Identify** the limiting reactant for a reaction and use it to calculate theoretical yield.
2. **Perform** calculations involving percentage yield.

Limiting Reactants and Theoretical Yield

To drive a car, you need gasoline in the tank and oxygen from the air. When the gasoline runs out, you can't go any farther even though there is still plenty of oxygen. In other words, the gasoline limits the distance you can travel because it runs out and the reaction in the engine stops.

In the previous section, you assumed that 100% of the reactants changed into products. And that is what should happen theoretically. But in the real world, other factors, such as the amounts of all reactants, the completeness of the reaction, and product lost in the process, can limit the yield of a reaction. The analogy of assembling homecoming mums for a fund raiser, as shown in **Figure 3,** will help you understand that whatever is in short supply will limit the quantity of product made.

Figure 3
The number of mums these students can assemble will be limited by the component that runs out first.

The Limiting Reactant Forms the Least Product

The students assembling mums use one helmet, one flower, eight blue ribbons, six white ribbons, and two bells to make each mum. As a result, the students cannot make any more mums once any one of these items is used up. Likewise, the reactants of a reaction are seldom present in ratios equal to the mole ratio in the balanced equation. So one of the reactants is used up first. For example, one way to to make H_2 is

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

If you combine 0.23 mol Zn and 0.60 mol HCl, would they react completely? Using the coefficients from the balanced equation, you can predict that 0.23 mol Zn can form 0.23 mol H_2, and 0.60 mol HCl can form 0.30 mol H_2. Zinc is called the **limiting reactant** because the zinc limits the amount of product that can form. The zinc is used up first by the reaction. The HCl is the **excess reactant** because there is more than enough HCl present to react with all of the Zn. There will be some HCl left over after the reaction stops.

Again, think of the mums, and look at **Figure 4.** The supplies at left are the available reactants. The products formed are the finished mums. The limiting reactant is the flowers because they are completely used up first. The ribbons, helmets, and bells are excess reactants because there are some of each of these items left over, at right. You can determine the limiting reactant by calculating the amount of product that each reactant could form. Whichever reactant would produce the least amount of product is the limiting reactant.

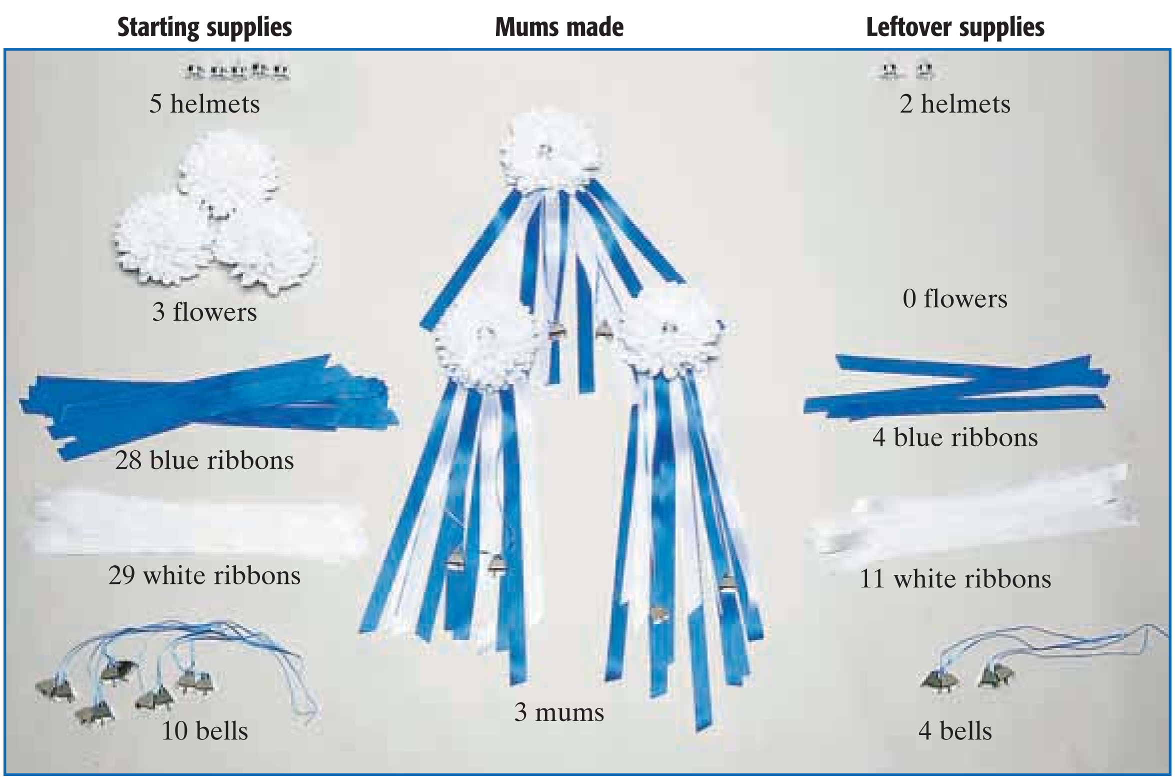
limiting reactant

the substance that controls the quantity of product that can form in a chemical reaction

excess reactant

the substance that is not used up completely in a reaction

Figure 4
The flowers are in short supply. They are the limiting reactant for assembling these homecoming mums.



Determine Theoretical Yield from the Limiting Reactant

So far you have done only calculations that assume reactions happen perfectly. The maximum quantity of product that a reaction could theoretically make if everything about the reaction works perfectly is called the *theoretical yield.* The theoretical yield of a reaction should always be calculated based on the limiting reactant.

In the reaction of Zn with HCl, the theoretical yield is 0.23 mol H_2 even though the HCl could make 0.30 mol H_2.

SAMPLE PROBLEM E

Limiting Reactants and Theoretical Yield

Identify the limiting reactant and the theoretical yield of phosphorous acid, H_3PO_3, if 225 g of PCl_3 is mixed with 123 g of H_2O.

$$PCl_3 + 3H_2O \longrightarrow H_3PO_3 + 3HCl$$

1 Gather information.

- mass PCl_3 = 225 g PCl_3
- mass H_2O = 123 g H_2O
- mass H_3PO_3 = ? g H_3PO_3
- molar mass PCl_3 = 137.32 g/mol
- molar mass H_2O = 18.02 g/mol
- molar mass H_3PO_3 = 82.00 g/mol
- From the balanced equation:
  1 mol PCl_3 = 1 mol H_3PO_3 and 3 mol H_2O = 1 mol H_3PO_3.

PRACTICE HINT

Whenever a problem gives you quantities of two or more reactants, you must determine the limiting reactant and use it to determine the theoretical yield.

2 Plan your work.

Set up problems that will calculate the mass of H_3PO_3 you would expect to form from each reactant.

3 Calculate.

$$? \text{ g } H_3PO_3 = 225 \cancel{\text{ g } PCl_3} \times \frac{1 \cancel{\text{ mol } PCl_3}}{137.32 \cancel{\text{ g } PCl_3}} \times \frac{1 \cancel{\text{ mol } H_3PO_3}}{1 \cancel{\text{ mol } PCl_3}} \times \frac{82.00 \text{ g } H_3PO_3}{1 \cancel{\text{ mol } H_3PO_3}} = 134 \text{ g } H_3PO_3$$

$$? \text{ g } H_3PO_3 = 123 \cancel{\text{ g } H_2O} \times \frac{1 \cancel{\text{ mol } H_2O}}{18.02 \cancel{\text{ g } H_2O}} \times \frac{1 \cancel{\text{ mol } H_3PO_3}}{3 \cancel{\text{ mol } H_2O}} \times \frac{82.00 \text{ g } H_3PO_3}{1 \cancel{\text{ mol } H_3PO_3}} = 187 \text{ g } H_3PO_3$$

PCl_3 is the limiting reactant. The theoretical yield is 134 g H_3PO_3.

4 Verify your result.

The units of the answer are correct, and estimating gives 128.

PRACTICE

Using the reaction above, identify the limiting reactant and the theoretical yield (in grams) of HCl for each pair of reactants.

1. 3.00 mol PCl_3 and 3.00 mol H_2O
2. 75.0 g PCl_3 and 75.0 g H_2O
3. 1.00 mol of PCl_3 and 50.0 g of H_2O

PROBLEM SOLVING SKILL

Limiting Reactants and the Food You Eat

In industry, the cheapest reactant is often used as the excess reactant. In this way, the expensive reactant is more completely used up. In addition to being cost-effective, this practice can be used to control which reactions happen. In the production of cider vinegar from apple juice, the apple juice is first kept where there is no oxygen so that the microorganisms in the juice break down the sugar, glucose, into ethanol and carbon dioxide. The resulting solution is hard cider.

Having excess oxygen in the next step allows the organisms to change ethanol into acetic acid, resulting in cider vinegar. Because the oxygen in the air is free and is easy to get, the makers of cider vinegar constantly pump air through hard cider as they make it into vinegar. Ethanol, which is not free, is the limiting reactant and is used up in the reaction.

Cost is also used to choose the excess reactant when making banana flavoring, isopentyl acetate. Acetic acid is the excess reactant because it costs much less than isopentyl alcohol.

$$CH_3COOH + C_5H_{11}OH \longrightarrow CH_3COOC_5H_{11} + H_2O$$
$$\text{acetic acid} + \text{isopentyl alcohol} \longrightarrow \text{isopentyl acetate} + \text{water}$$

As shown in **Figure 5,** when compared mole for mole, isopentyl alcohol is more than twice as expensive as acetic acid. When a large excess of acetic acid is present, almost all of the isopentyl alcohol reacts.

Choosing the excess and limiting reactants based on cost is also helpful in areas outside of chemistry. In making the homecoming mums, the flower itself is more expensive than any of the other materials, so it makes sense to have an excess of ribbons and charms. The expensive flowers are the limiting reactant.

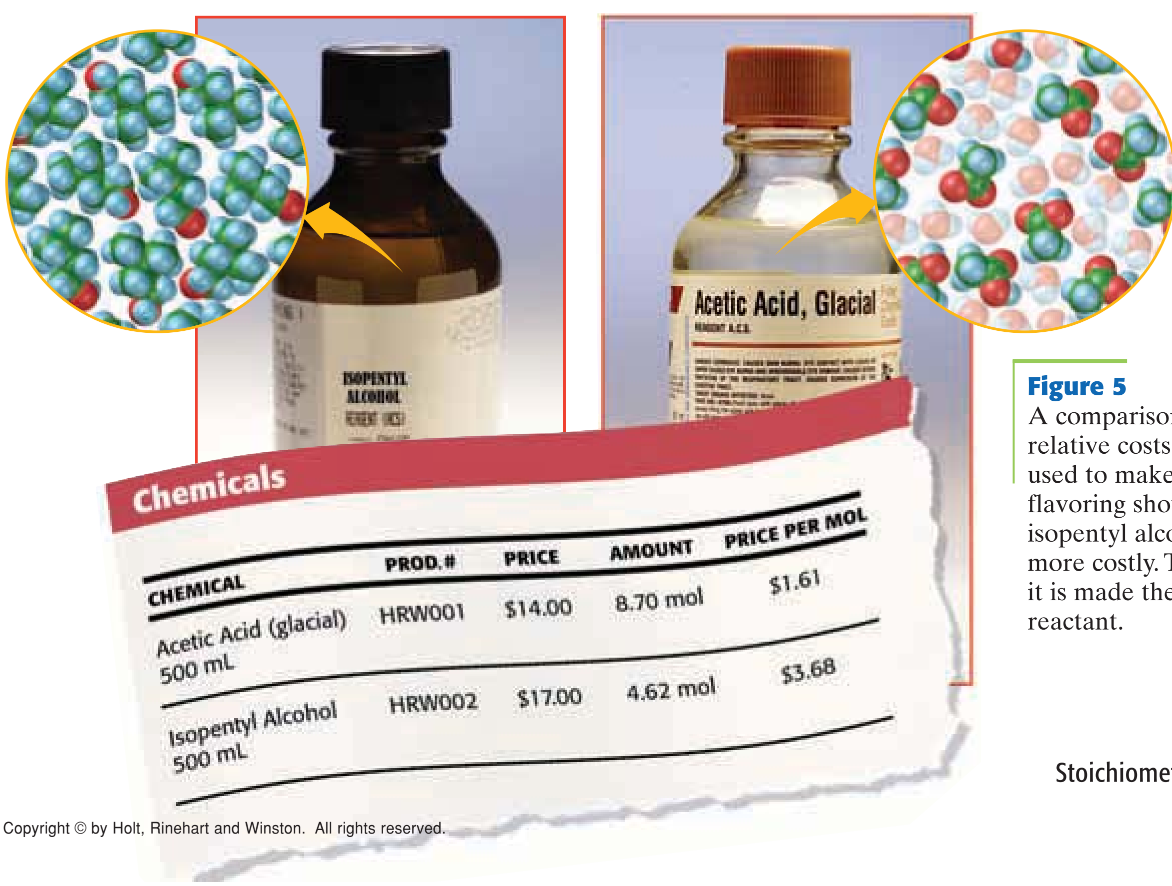


Figure 5
A comparison of the relative costs of chemicals used to make banana flavoring shows that isopentyl alcohol is more costly. That is why it is made the limiting reactant.

Table 1 Predictions and Results for Isopentyl Acetate Synthesis

Reactants	Formula	Mass present	Amount present	Amount left over
Isopentyl alcohol	$C_5H_{11}OH$	500.0 g	5.67 mol (limiting reactant)	0.0 mol
Acetic acid	CH_3COOH	1.25×10^3 g	20.8 mol	15.1 mol
Products	**Formula**	**Amount expected**	**Theoretical yield (mass expected)**	**Actual yield (mass produced)**
Isopentyl acetate	$CH_3COOC_5H_{11}$	5.67 mol	738 g	591 g
Water	H_2O	5.67 mol	102 g	81.6 g

Actual Yield and Percentage Yield

actual yield

the measured amount of a product of a reaction

Although equations tell you what *should* happen in a reaction, they cannot always tell you what *will* happen. For example, sometimes reactions do not make all of the product predicted by stoichiometric calculations, or the theoretical yield. In most cases, the **actual yield,** the mass of product actually formed, is less than expected. Imagine that a worker at the flavoring factory mixes 500.0 g isopentyl alcohol with 1.25×10^3 g acetic acid. The actual and theoretical yields are summarized in **Table 1.** Notice that the actual yield is less than the mass that was expected.

There are several reasons why the actual yield is usually less than the theoretical yield in chemical reactions. Many reactions do not completely use up the limiting reactant. Instead, some of the products turn back into reactants so that the final result is a mixture of reactants and products. In many cases the main product must go through additional steps to purify or separate it from other chemicals. For example, banana flavoring must be *distilled,* or isolated based on its boiling point. Solid compounds, such as sugar, must be recrystallized. Some of the product may be lost in the process. There also may be other reactions, called *side reactions*, that can use up reactants without making the desired product.

Determining Percentage Yield

The ratio relating the actual yield of a reaction to its theoretical yield is called the *percentage yield* and describes the efficiency of a reaction. Calculating a percentage yield is similar to calculating a batting average. A batter might get a hit every time he or she is at bat. This is the "theoretical yield." But no player has gotten a hit every time. Suppose a batter gets 41 hits in 126 times at bat. The batting average is 41 (the actual hits) divided by 126 (the possible hits theoretically), or 0.325. In the example described in **Table 1,** the theoretical yield for the reaction is 738 g. The actual yield is 591 g. The percentage yield is

$$\text{percentage yield} = \frac{591 \text{ g (actual yield)}}{738 \text{ g (theoretical yield)}} \times 100 = 80.1\%$$

SAMPLE PROBLEM F

Calculating Percentage Yield

Determine the limiting reactant, the theoretical yield, and the percentage yield if 14.0 g N_2 are mixed with 9.0 g H_2, and 16.1 g NH_3 form.

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

1 Gather information.

- mass N_2 = 14.0 g N_2
- mass H_2 = 9.0 g H_2
- theoretical yield of NH_3 = ? g NH_3
- actual yield of NH_3 = 16.1 g NH_3
- From the balanced equation: 1 mol N_2 = 2 mol NH_3 and 3 mol H_2 = 2 mol NH_3.
- molar mass N_2 = 28.02 g/mol
- molar mass H_2 = 2.02 g/mol
- molar mass NH_3 = 17.04 g/mol

2 Plan your work.

Set up problems that will calculate the mass of NH_3 you would expect to form from each reactant.

3 Calculate.

$$? \text{ g } NH_3 = 14.0 \cancel{\text{ g } N_2} \times \frac{1 \cancel{\text{ mol } N_2}}{28.02 \cancel{\text{ g } N_2}} \times \frac{2 \cancel{\text{ mol } NH_3}}{1 \cancel{\text{ mol } N_2}} \times \frac{17.04 \text{ g } NH_3}{1 \cancel{\text{ mol } NH_3}} = 17.0 \text{ g } NH_3$$

$$? \text{ g } NH_3 = 9.0 \cancel{\text{ g } H_2} \times \frac{1 \cancel{\text{ mol } H_2}}{2.02 \cancel{\text{ g } H_2}} \times \frac{2 \cancel{\text{ mol } NH_3}}{3 \cancel{\text{ mol } H_2}} \times \frac{17.04 \text{ g } NH_3}{1 \cancel{\text{ mol } NH_3}} = 51 \text{ g } NH_3$$

- The smaller quantity made, 17.0 g NH_3, is the theoretical yield so the limiting reactant is N_2.
- The percentage yield is calculated:

$$\text{percentage yield} = \frac{16.1 \text{ g (actual yield)}}{17.0 \text{ g (theoretical yield)}} \times 100 = 94.7\%$$

4 Verify your result.

The units of the answer are correct. The percentage yield is less than 100%, so the final calculation is probably set up correctly.

PRACTICE HINT

If an amount of product actually formed is given in a problem, this is the reaction's actual yield.

PRACTICE

Determine the limiting reactant and the percentage yield for each of the following.

1. 14.0 g N_2 react with 3.15 g H_2 to give an actual yield of 14.5 g NH_3.
2. In a reaction to make ethyl acetate, 25.5 g CH_3COOH react with 11.5 g C_2H_5OH to give a yield of 17.6 g $CH_3COOC_2H_5$.

$$CH_3COOH + C_2H_5OH \longrightarrow CH_3COOC_2H_5 + H_2O$$

3. 16.1 g of bromine are mixed with 8.42 g of chlorine to give an actual yield of 21.1 g of bromine monochloride.

PROBLEM SOLVING SKILL

Determining Actual Yield

Although the actual yield can only be determined experimentally, a close estimate can be calculated if the percentage yield for a reaction is known. The percentage yield in a particular reaction is usually fairly consistent. For example, suppose an industrial chemist determined the percentage yield for six tries at making banana flavoring and found the results were 80.0%, 82.1%, 79.5%, 78.8%, 80.5%, and 81.9%. In the future, the chemist can expect a yield of around 80.5%, or the average of these results.

If the chemist has enough isopentyl alcohol to make 594 g of the banana flavoring theoretically, then an actual yield of around 80.5% of that, or 478 g, can be expected.

SAMPLE PROBLEM G

Calculating Actual Yield

How many grams of $CH_3COOC_5H_{11}$ should form if 4808 g are theoretically possible and the percentage yield for the reaction is 80.5%?

1 Gather information.

- theoretical yield of $CH_3COOC_5H_{11}$ = 4808 g $CH_3COOC_5H_{11}$
- actual yield of $CH_3COOC_5H_{11}$ = ? g $CH_3COOC_5H_{11}$
- percentage yield = 80.5%

2 Plan your work.

Use the percentage yield and the theoretical yield to calculate the actual yield expected.

3 Calculate.

$$80.5\% = \frac{\text{actual yield}}{4808 \text{ g}} \times 100$$

actual yield = 4808 g × 0.805 = 3.87×10^3 g $CH_3COOC_5H_{11}$

4 Verify your result.

The units of the answer are correct. The actual yield is less than the theoretical yield, as it should be.

PRACTICE HINT

The actual yield should always be less than the theoretical yield. A wrong answer that is greater than the theoretical yield can result if you accidentally reverse the actual and theoretical yields.

PRACTICE

PROBLEM SOLVING SKILL

1. The percentage yield of NH_3 from the following reaction is 85.0%. What actual yield is expected from the reaction of 1.00 kg N_2 with 225 g H_2?

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

2. If the percentage yield is 92.0%, how many grams of CH_3OH can be made by the reaction of 5.6×10^3 g CO with 1.0×10^3 g H_2?

$$CO + 2H_2 \longrightarrow CH_3OH$$

3. Suppose that the percentage yield of BrCl is 90.0%. How much BrCl can be made by reacting 338 g Br_2 with 177 g Cl_2?

2 Section Review

UNDERSTANDING KEY IDEAS

1. Distinguish between limiting reactant and excess reactant in a chemical reaction.
2. How do manufacturers decide which reactant to use in excess in a chemical reaction?
3. How do you calculate the percentage yield of a chemical reaction?
4. Give two reasons why a 100% yield is not obtained in actual chemical manufacturing processes.
5. How do the values of the theoretical and actual yields generally compare?

PRACTICE PROBLEMS

6. A chemist reacts 8.85 g of iron with an excess of hydrogen chloride to form hydrogen gas and iron(II) chloride. Calculate the theoretical yield and the percentage yield of hydrogen if 0.27 g H_2 are collected.
7. Use the chemical reaction below to answer the questions that follow.

$$P_4O_{10} + H_2O \longrightarrow H_3PO_4$$

   a. Balance the equation.
   b. Calculate the theoretical yield if 100.0 g P_4O_{10} react with 200.0 g H_2O.
   c. If the actual mass recovered is 126.2 g H_3PO_4, what is the percentage yield?
8. Titanium dioxide is used as a white pigment in paints. If 3.5 mol $TiCl_4$ reacts with 4.5 mol O_2, which is the limiting reactant? How many moles of each product are produced? How many moles of the excess reactant remain?

$$TiCl_4 + O_2 \longrightarrow TiO_2 + 2Cl_2$$

9. If 1.85 g Al reacts with an excess of copper(II) sulfate and the percentage yield of Cu is 56.6%, what mass of Cu is produced?
10. Quicklime, CaO, can be prepared by roasting limestone, $CaCO_3$, according to the chemical equation below. When 2.00×10^3 g of $CaCO_3$ are heated, the actual yield of CaO is 1.05×10^3 g. What is the percentage yield?

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

11. Magnesium powder reacts with steam to form magnesium hydroxide and hydrogen gas.
   a. Write a balanced equation for this reaction.
   b. What is the percentage yield if 10.1 g Mg reacts with an excess of water and 21.0 g $Mg(OH)_2$ is recovered?
   c. If 24 g Mg is used and the percentage yield is 95%, how many grams of magnesium hydroxide should be recovered?
12. Use the chemical reaction below to answer the questions that follow.

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(g)$$

   a. What is the limiting reactant when 19.9 g CuO react with 2.02 g H_2?
   b. The actual yield of copper was 15.0 g. What is the percentage yield?
   c. How many grams of Cu can be collected if 20.6 g CuO react with an excess of hydrogen with a yield of 91.0%?

CRITICAL THINKING

13. A chemist reacts 20 mol H_2 with 20 mol O_2 to produce water. Assuming all of the limiting reactant is converted to water in the reaction, calculate the amount of each substance present after the reaction.
14. A pair of students performs an experiment in which they collect 27 g CaO from the decomposition of 41 g $CaCO_3$. Are these results reasonable? Explain your answer using percentage yield.

SECTION 3

Stoichiometry and Cars

OBJECTIVES

1. **Relate** volume calculations in stoichiometry to the inflation of automobile safety air bags.
2. **Use** the concept of limiting reactants to explain why fuel-air ratios affect engine performance.
3. **Compare** the efficiency of pollution-control mechanisms in cars using percentage yield.

Stoichiometry and Safety Air Bags

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So far you have examined stoichiometry in a number of chemical reactions, including making banana flavoring and ammonia. Now it is time to look at stoichiometry in terms of something a little more familiar—a car. Stoichiometry is important in many aspects of automobile operation and safety. First, let's look at how stoichiometry can help keep you safe should you ever be in an accident. Air bags have saved the lives of many people involved in accidents. And the design of air bags requires an understanding of stoichiometry.

An Air Bag Could Save Your Life

Air bags are designed to protect people in a car from being hurt during a high-speed collision. When inflated, air bags slow the motion of a person so that he or she does not strike the steering wheel, windshield, or dashboard with as much force.

Stoichiometry is used by air-bag designers to ensure that air bags do not underinflate or overinflate. Bags that underinflate do not provide enough protection, and bags that overinflate can cause injury by bouncing the person back with too much force. Therefore, the chemicals must be present in just the right proportions. To protect riders, air bags must inflate within one-tenth of a second after impact. The basic components of most systems that make an air bag work are shown in **Figure 6.** A front-end collision transfers energy to a crash sensor that causes an igniter to fire. The igniter provides the energy needed to start a very fast reaction that produces gas in a mixture called the *gas generant*. The igniter also raises the temperature and pressure within the inflator (a metal vessel) so that the reaction happens fast enough to fill the bag before the rider strikes it. A high-efficiency filter keeps the hot reactants and the solid products away from the rider, and additional chemicals are used to make the products safer.

Air-Bag Design Depends on Stoichiometric Precision

The materials used in air bags are constantly being improved to make air bags safer and more effective. Many different materials are used. One of the first gas generants used in air bags is still in use in some systems. It is a solid mixture of sodium azide, NaN_3, and an oxidizer. The gas that inflates the bag is almost pure nitrogen gas, N_2, which is produced in the following decomposition reaction.

$$2NaN_3(s) \longrightarrow 2Na(s) + 3N_2(g)$$

However, this reaction does not inflate the bag enough, and the sodium metal is dangerously reactive. Oxidizers such as ferric oxide, Fe_2O_3, are included, which react rapidly with the sodium. Energy is released, which heats the gas and causes the gas to expand and fill the bag.

$$6Na(s) + Fe_2O_3(s) \longrightarrow 3Na_2O(s) + 2Fe(s) + \text{energy}$$

One product, sodium oxide, Na_2O, is extremely corrosive. Water vapor and CO_2 from the air react with it to form less harmful $NaHCO_3$.

$$Na_2O(s) + 2CO_2(g) + H_2O(g) \longrightarrow 2NaHCO_3(s)$$

The mass of gas needed to fill an air bag depends on the density of the gas. Gas density depends on temperature. To find the amount of gas generant to put into each system, designers must know the stoichiometry of the reactions and account for changes in temperature and thus the density of the gas.

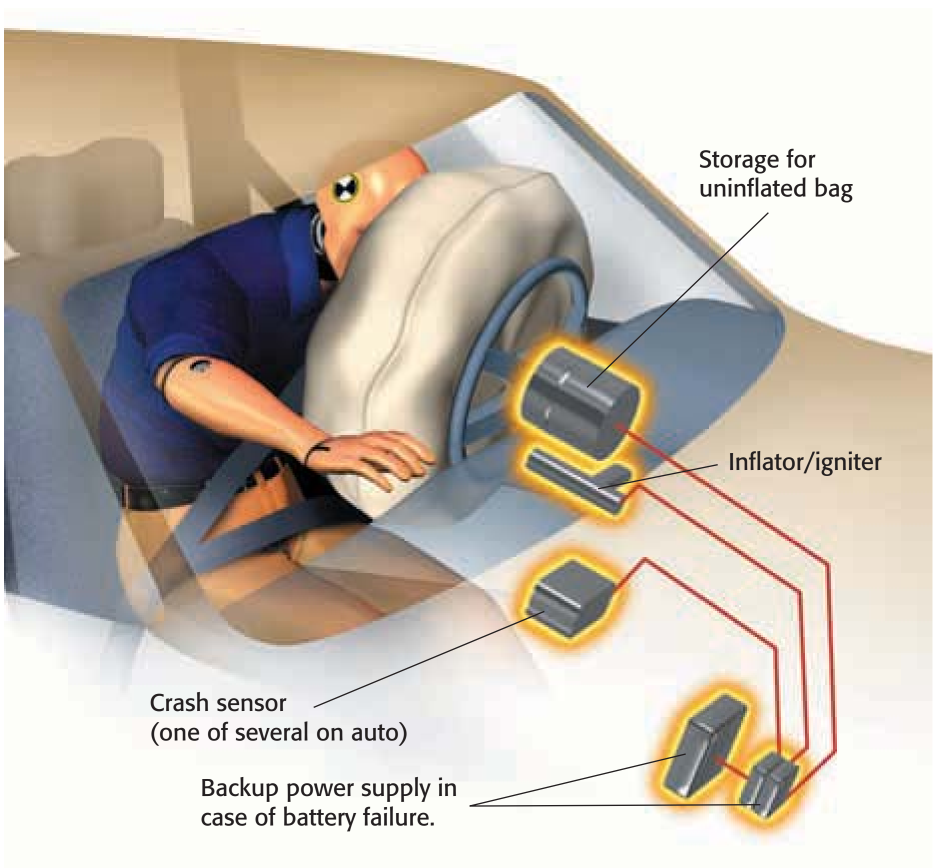


Figure 6
Inflating an air bag requires a rapid series of events, eventually producing nitrogen gas to inflate the air bag.

SAMPLE PROBLEM H

Air-Bag Stoichiometry

Assume that 65.1 L N_2 inflates an air bag to the proper size. What mass of NaN_3 must be used? (density of N_2 = 0.92 g/L)

1 Gather information.

- Write a balanced chemical equation

$$2NaN_3(s) \longrightarrow 2Na(s) + 3N_2(g)$$

- volume of N_2 = 65.1 L N_2
- density of N_2 = 0.92 g/L
- molar mass of N_2 = 28.02 g/mol
- mass of reactant = ? g NaN_3
- molar mass of NaN_3 = 65.02 g/mol
- From the balanced equation: 2 mol NaN_3 = 3 mol N_2.

2 Plan your work.

Start with the volume of N_2, and change it to moles using density and molar mass. Then use the mole ratio followed by the molar mass of NaN_3.

3 Calculate.

$$? \text{ g } NaN_3 = 65.1 \text{ } \cancel{\text{L } N_2} \times \frac{0.92 \text{ } \cancel{\text{g } N_2}}{1 \text{ } \cancel{\text{L } N_2}} \times \frac{1 \text{ } \cancel{\text{mol } N_2}}{28.02 \text{ } \cancel{\text{g } N_2}} \times$$

$$\frac{2 \text{ } \cancel{\text{mol } NaN_3}}{3 \text{ } \cancel{\text{mol } N_2}} \times \frac{65.02 \text{ g } NaN_3}{1 \text{ } \cancel{\text{mol } NaN_3}} = 93 \text{ g } NaN_3$$

4 Verify your result.

The number of significant figures is correct. Estimating gives 90.

PRACTICE HINT

Gases are measured by volume, just as liquids are. In problems with volume, you can use the density to convert to mass and the molar mass to convert to moles. Then use the mole ratio, just as in any other stoichiometry problem.

PRACTICE

PROBLEM SOLVING SKILL

1. How many grams of Na form when 93 g NaN_3 react?
2. The Na formed during the breakdown of NaN_3 reacts with Fe_2O_3. How many grams of Fe_2O_3 are needed to react with 35.3 g Na?

$$6Na(s) + Fe_2O_3(s) \longrightarrow 3Na_2O(s) + 2Fe(s)$$

3. The Na_2O formed in the above reaction is made less harmful by the reaction below. How many grams of $NaHCO_3$ are made from 44.7 g Na_2O?

$$Na_2O(s) + 2CO_2(g) + H_2O(g) \longrightarrow 2NaHCO_3(s)$$

4. Suppose the reaction below was used to fill a 65.1 L air bag with CO_2 and the density of CO_2 at the air bag temperature is 1.35 g/L.

$$NaHCO_3 + HC_2H_3O_2 \longrightarrow NaC_2H_3O_2 + CO_2 + H_2O$$

a. How many grams of $NaHCO_3$ are needed?

b. How many grams of $HC_2H_3O_2$ are needed?

Stoichiometry and Engine Efficiency

The efficiency of a car's engine depends on having the correct stoichiometric ratio of gasoline and oxygen. Although gasoline used in automobiles is a mixture, it can be treated as if it were pure isooctane, one of the many compounds whose formula is C_8H_{18}. (This compound has a molar mass that is about the same as the weighted average of the compounds in actual gasoline.) The other reactant in gasoline combustion is oxygen, which is about 21% of air by volume. The reaction for gasoline combustion can be written as follows.

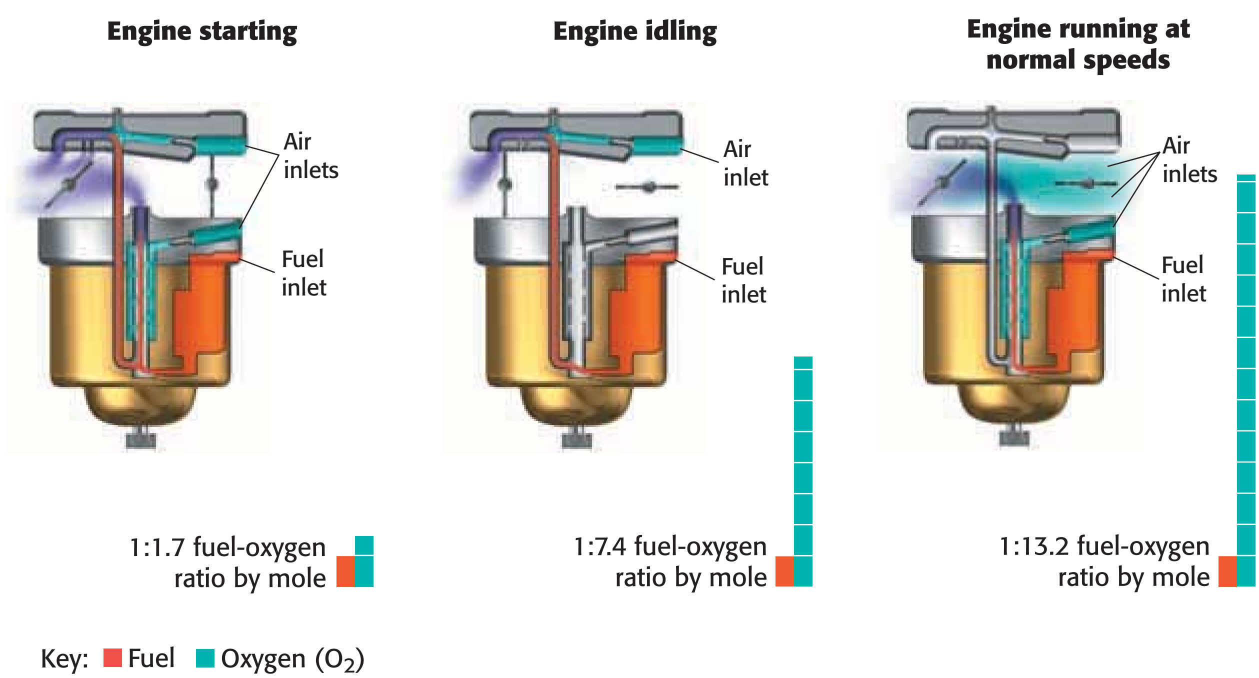
$$2C_8H_{18}(g) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(g)$$

Engine Efficiency Depends on Reactant Proportions

For efficient combustion, the above two reactants must be mixed in a mole ratio that is close to the one shown in the balanced chemical equation, that is 2:25, or 1:12.5. If there is not enough of either reactant, the engine might stall. For example, if you pump the gas pedal too many times before starting, the mixture of reactants in the engine will contain an excess of gasoline, and the lack of oxygen may prevent the mixture from igniting. This is referred to as "flooding the engine." On the other hand, if there is too much oxygen and not enough gasoline, the engine will stall just as if the car were out of gas.

Although the best stoichiometric mixture of fuel and oxygen is 1:12.5 in terms of moles, this is not the best mixture to use all the time. **Figure 7** shows a model of a carburetor controlling the fuel-oxygen ratio in an engine that is starting, idling, and running at normal speeds. Carburetors are often used in smaller engines, such as those in lawn mowers. Computer-controlled fuel injectors have taken the place of carburetors in car engines.

Figure 7
The fuel-oxygen ratio changes depending on what the engine is doing.



SAMPLE PROBLEM I

Air-Fuel Ratio

A cylinder in a car's engine draws in 0.500 L of air. How many milliliters of liquid isooctane should be injected into the cylinder to completely react with the oxygen present? The density of isooctane is 0.692 g/mL, and the density of oxygen is 1.33 g/L. Air is 21% oxygen by volume.

1 Gather information.

- Write a balanced equation for the chemical reaction.

$$2C_8H_{18} + 25O_2 \longrightarrow 16CO_2 + 18H_2O$$

- volume of air = 0.500 L air
- percentage of oxygen in air: 21% by volume
- Organize the data in a table.

Reactant	Formula	Molar mass	Density	Volume
Oxygen	O_2	32.00 g/mol	1.33 g/L	? L
Isooctane	C_8H_{18}	114.26 g/mol	0.692 g/mL	? mL

- From the balanced equation: 2 mol C_8H_{18} = 25 mol O_2.

2 Plan your work.

Use the percentage by volume of O_2 in air to find the volume of O_2. Then set up a volume-volume problem.

3 Calculate.

$$? \text{ mL } C_8H_{18} = 0.500 \cancel{\text{L air}} \times \frac{21 \cancel{\text{L } O_2}}{100 \cancel{\text{L air}}} \times \frac{1.33 \cancel{\text{g } O_2}}{1 \cancel{\text{L } O_2}} \times \frac{1 \cancel{\text{mol } O_2}}{32.00 \cancel{\text{g } O_2}} \times$$

$$\frac{2 \cancel{\text{mol } C_8H_{18}}}{25 \cancel{\text{mol } O_2}} \times \frac{114.26 \cancel{\text{g } C_8H_{18}}}{1 \cancel{\text{mol } C_8H_{18}}} \times \frac{1 \text{ mL } C_8H_{18}}{0.692 \cancel{\text{g } C_8H_{18}}} = 5.76 \times 10^{-2} \text{ mL } C_8H_{18}$$

4 Verify your result.

The denominator is about 10 times larger than the numerator, so the answer in mL should be about one-tenth of the original volume in L.

PRACTICE HINT

Remember that in problems with volumes, you must be sure that the volume unit in the density matches the volume unit given or wanted.

PRACTICE

PROBLEM SOLVING SKILL

1. A V-8 engine has eight cylinders each having a 5.00×10^2 cm^3 capacity. How many cycles are needed to completely burn 1.00 mL of isooctane? (One cycle is the firing of all eight cylinders.)
2. How many milliliters of isooctane are burned during 25.0 cycles of a V-6 engine having six cylinders each having a 5.00×10^2 cm^3 capacity?
3. Methyl alcohol, CH_3OH, with a density of 0.79 g/mL, can be used as fuel in race cars. Calculate the volume of air needed for the complete combustion of 51.0 mL CH_3OH.

Table 2 Clean Air Act Standards for 1996 Air Pollution

Pollutant	Cars	Light trucks	Motorcycles
Hydrocarbons	0.25 g/km	0.50 g/km	5.0 g/km
Carbon monoxide	2.1 g/km	2.1–3.1 g/km, depending on truck size	12 g/km
Oxides of nitrogen (NO, NO_2)	0.25 g/km	0.25–0.68 g/km, depending on truck size	not regulated

Stoichioimetry and Pollution Control

Automobiles are the primary source of air pollution in many parts of the world. The Clean Air Act was enacted in 1968 to address the issue of smog and other forms of pollution caused by automobile exhaust. This act has been amended to set new, more restrictive emission-control standards for automobiles driven in the United States. **Table 2** lists the standards for pollutants in exhaust set in 1996 by the U.S. Environmental Protection Agency.

The Fuel-Air Ratio Influences the Pollutants Formed

The equation for the combustion of "isooctane" shows most of what happens when gasoline burns, but it does not tell the whole story. For example, if the fuel-air mixture does not have enough oxygen, some carbon monoxide will be produced instead of carbon dioxide. When a car is started, there is less air, so fairly large amounts of carbon monoxide are formed, and some unburned fuel (hydrocarbons) also comes out in the exhaust. In cold weather, an engine needs more fuel to start, so larger amounts of unburned hydrocarbons and carbon monoxide come out as exhaust. These hydrocarbons are involved in forming smog. So the fuel-air ratio is a key factor in determining how much pollution forms.

Another factor in auto pollution is the reaction of nitrogen and oxygen at the high temperatures inside the engine to form small amounts of highly reactive nitrogen oxides, including NO and NO_2.

$$N_2(g) + O_2(g) \longrightarrow 2NO(g)$$

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

One of the Clean Air Act standards limits the amount of nitrogen oxides that a car can emit. These compounds react with oxygen to form another harmful chemical, ozone, O_3.

$$NO_2(g) + O_2(g) \longrightarrow 2NO(g) + O_3(g)$$

Because these reactions are started by energy from the sun's ultraviolet light, they form what is referred to as *photochemical smog*. The harmful effects of photochemical smog are caused by very small concentrations of pollutants, including unburned hydrocarbon fuel.

internet connect

www.scilinks.org
Topic: Air Pollution
SciLinks code: HW4133

SCILINKS Maintained by the National Science Teachers Association

Meeting the Legal Limits Using Stoichiometry

Automobile manufacturers use stoichiometry to predict when adjustments will be necessary to keep exhaust emissions within legal limits. Because the units in **Table 2** are *grams per kilometer*, auto manufacturers must consider how much fuel the vehicle will burn to move a certain distance. Automobiles with better gas mileage will use less fuel per kilometer, resulting in lower emissions per kilometer.

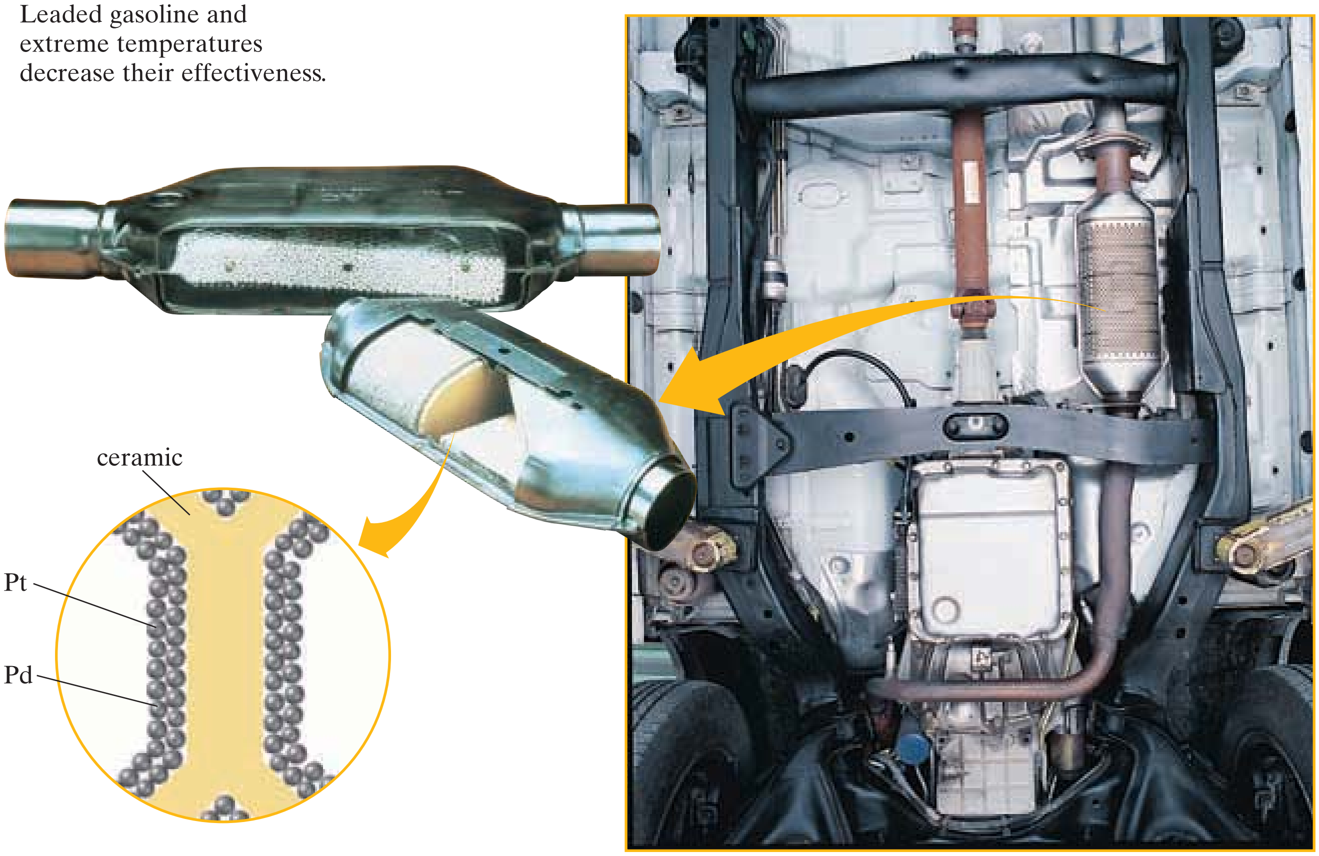
internet connect
www.scilinks.org
Topic: Cataytic Converters
SciLinks code: HW4026
SciLinks Maintained by the National Science Teachers Association

Catalytic Converters Can Help

All cars that are currently manufactured in the United States are built with catalytic converters, like the one shown in **Figure 8,** to treat the exhaust gases before they are released into the air. Platinum, palladium, or rhodium in these converters act as catalysts and increase the rate of the decomposition of NO and of NO_2 into N_2 and O_2, harmless gases already found in the atmosphere. Catalytic converters also speed the change of CO into CO_2 and the change of unburned hydrocarbons into CO_2 and H_2O. These hydrocarbons are involved in the formation of ozone and smog, so it is important that unburned fuel does not come out in the exhaust.

Catalytic converters perform at their best when the exhaust gases are hot and when the ratio of fuel to air in the engine is very close to the proper stoichiometric ratio. Newer cars include on-board computers and oxygen sensors to make sure the proper fuel-air ratio is automatically maintained, so that the engine and the catalytic converter work at top efficiency.

Figure 8
Catalytic converters are used to decrease nitrogen oxides, carbon monoxide, and hydrocarbons in exhaust. Leaded gasoline and extreme temperatures decrease their effectiveness.



SAMPLE PROBLEM J

Calculating Yields: Pollution

What mass of ozone, O_3, can be produced from 3.50 g of NO_2 contained in a car's exhaust? The equation is as follows.

$$NO_2(g) + O_2(g) \longrightarrow NO(g) + O_3(g)$$

1 Gather information.

- mass of NO_2 = 3.50 g NO_2
- molar mass of NO_2 = 46.01 g/mol
- mass of O_3 = ? g O_3
- molar mass of O_3 = 48.00 g/mol
- From the balanced equation: 1 mol NO_2 = 1 mol O_3.

2 Plan your work.

This is a mass-mass problem.

3 Calculate.

$$? \text{ g } O_3 = 3.50 \cancel{\text{ g } NO_2} \times \frac{1 \cancel{\text{ mol } NO_2}}{46.01 \cancel{\text{ g } NO_2}} \times \frac{1 \cancel{\text{ mol } O_3}}{1 \cancel{\text{ mol } NO_2}} \times \frac{48.00 \text{ g } O_3}{1 \cancel{\text{ mol } O_3}} = 3.65 \text{ g } O_3$$

4 Verify your result.

The denominator and numerator are almost equal, so the mass of product is almost the same as the mass of reactant.

PRACTICE HINT

This is a review of the first type of stoichiometric calculation that you learned.

PRACTICE

1 A catalytic converter combines 2.55 g CO with excess O_2. What mass of CO_2 forms?

PROBLEM SOLVING SKILL

3 Section Review

UNDERSTANDING KEY IDEAS

1. What is the main gas in an air bag that is inflated using the NaN_3 reaction?
2. How do you know that the correct mole ratio of isooctane to oxygen is 1:12.5?
3. What do the catalysts in the catalytic converters accomplish?
4. Give at least two results of too little air being in a running engine.

PRACTICE PROBLEMS

5. Assume that 22.4 g of NaN_3 react completely in an air bag. What mass of Na_2O is produced after complete reaction of the Na with Fe_2O_3?
6. Na_2O eventually reacts with CO_2 and H_2O to form $NaHCO_3$. What mass of $NaHCO_3$ is formed when 44.4 g Na_2O completely react?

CRITICAL THINKING

7. Why are nitrogen oxides in car exhaust, even though there is no nitrogen in the fuel?
8. Why not use the following reaction to produce N_2 in an air bag?

$$NH_3(g) + O_2(g) \longrightarrow N_2(g) + H_2O(g)$$

9. Just after an automobile is started, you see water dripping off the end of the tail pipe. Is this normal? Why or why not?

9 CHAPTER HIGHLIGHTS

KEY IDEAS

SECTION ONE Calculating Quantities in Reactions

- Reaction stoichiometry compares the amounts of substances in a chemical reaction.
- Stoichiometry problems involving reactions can always be solved using mole ratios.
- Stoichiometry problems can be solved using three basic steps. First, change what you are given into moles. Second, use a mole ratio based on a balanced chemical equation. Third, change to the units needed for the answer.

SECTION TWO Limiting Reactants and Percentage Yield

- The limiting reactant is a reactant that is consumed completely in a reaction.
- The theoretical yield is the amount of product that can be formed from a given amount of limiting reactant.
- The actual yield is the amount of product collected from a real reaction.
- Percentage yield is the actual yield divided by the theoretical yield multiplied by 100. It is a measure of the efficiency of a reaction.

SECTION THREE Stoichiometry and Cars

- Stoichiometry is used in designing air bags for passenger safety.
- Stoichiometry is used to maximize a car's fuel efficiency.
- Stoichiometry is used to minimize the pollution coming from the exhaust of an auto.

KEY TERMS

stoichiometry

limiting reactant
excess reactant
actual yield

KEY SKILLS

Using Mole Ratios
Skills Toolkit 1 p. 303
Sample Problem A p. 304

Solving Stoichiometry Problems
Skills Toolkit 2 p. 305

Problems Involving Mass
Skills Toolkit 3 p. 306
Sample Problem B p. 307

Problems Involving Volume
Skills Toolkit 4 p. 308
Sample Problem C p. 309

Problems Involving Particles
Skills Toolkit 5 p. 310
Sample Problem D p. 310

Limiting Reactants and Theoretical Yield
Sample Problem E p. 314

Calculating Percentage Yield
Sample Problem F p. 317

Calculating Actual Yield
Sample Problem G p. 318

Air-Bag Stoichiometry
Sample Problem H p. 322

Air-Fuel Ratio
Sample Problem I p. 324

Calculating Yields: Pollution
Sample Problem J p. 327

Physical Setting/Chemistry
REGENTS EXAM PRACTICE

9

PART A

For each item, write on a separate piece of paper the number of the word, expression, or statement that best answers the item.

1. Which is most likely to be used to solve a stoichiometry problem?
(1) table of bond energies
(2) mole ratio from a balanced chemical equation
(3) chart of electron configurations
(4) a periodic table

2. Which procedure is most closely related to stoichiometry?
(1) measuring quantities of reactants and products
(2) studying reaction conditions
(3) determining empirical formulas
(4) measuring energy changes

3. What information is least likely to be used to solve mass-mass stoichiometry problems?
(1) coefficients used in a balanced chemical equation
(2) mole ratio for reactants and products
(3) table of electron configurations for each atom involved in the reaction
(4) table of atomic masses for each atom involved in the reaction

4. In the balanced chemical equation $A + 2B \longrightarrow C + 4D$, if you know the mass of A you can determine
(1) the mass of the other reactant, B, and the masses of the products, C and D.
(2) only the mass of C and D combined.
(3) only the mass of B.
(4) only the mass of A and B combined.

5. Which term is applied to the maximum amount of a product that can be obtained in a chemical reaction?
(1) actual yield
(2) mole ratio
(3) percentage yield
(4) theoretical yield

6. Which is most likely to be determined using the mole ratio of the reactants and products in a chemical reaction?
(1) energy that is either absorbed or released in the reaction
(2) mass of a product produced from a known mass of a reactant
(3) rate of the reaction
(4) chemical names of the reactants and products

7. What is the main concern to a chemist who is investigating a problem in stoichiometry?
(1) type of bonds found in compounds
(2) energy changes occurring in chemical reactions
(3) mass relationship in chemical reactions
(4) rate with which chemical reactions occur

8. What is the mole ratio of CO_2 to $C_6H_{12}O_6$ in the combustion reaction shown below?

$$C_6H_{12}O_6(s) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(g)$$

(1) 1:2
(2) 1:1
(3) 6:1
(4) 1:4

Regents Test-Taking Tip

Choose an answer to a question based on both what you already know as well as any information presented in the question.

9. What name is given to the quantity of reactant remaining after all the limiting reactant is completely consumed?
(1) product
(2) controlling reactant
(3) excess reactant
(4) catalyst

10. Identify the equation that can be used to solve a stoichiometry problem.
(1) $H_2(g) + O_2(g) \longrightarrow H_2O(g)$ + energy
(2) $H_2(g) + 2O_2(g) \longrightarrow 2H_2O(g)$ + energy
(3) $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$ + energy
(4) $2H_2(g) + 2O_2(g) \longrightarrow H_2O(g)$ + energy

11. What is the mole ratio of reactants O_2 to C_6H_{12} in the combustion reaction below?

$$C_6H_{12}(g) + 9O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(g)$$

(1) 1:2
(2) 9:1
(3) 1:1
(4) 1:9

12. For the reaction below, what is the mole ratio of $N_2(g)$ to $NH_3(g)$?

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

(1) 2:1
(2) 3:2
(3) 1:1
(4) 1:2

13. What name is given to the value calculated using the definition below?

$$\frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

(1) percent yield
(2) percent excess reactant
(3) percent limiting reactant
(4) percent error

14. For the reaction below, how many moles of a reactant are required to produce 18 mol NO_2?

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

(1) 9 mol O_2
(2) 18 mol O_2
(3) 2 mol NO
(4) 9 mol NO

15. Examine the following reaction.

$$SiO_2(s) + 4HF(g) \longrightarrow SiF_4(g) + 2H_2O(l)$$

When the reaction has gone to completion and all the $HF(g)$ is consumed, some $SiO_2(s)$ remains. The remaining quantity of $SiO_2(s)$ is called the
(1) limiting product.
(2) excess product.
(3) limiting reactant.
(4) excess reactant.

16. A student has calculated that a reaction can produce 10 g of a product as the maximum amount. When the reaction is performed, 9.5 g of the product is obtained. What is the percent yield of this reaction?
(1) 0.5%
(2) 5 %
(3) 9.5%
(4) 95%

17. Examine the following reaction.

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

When the reaction has gone to completion, all the $N_2(g)$ is consumed yet some $H_2(g)$ still remains. The quantity of $N_2(g)$ is called the
(1) limiting product.
(2) excess product.
(3) limiting reactant.
(4) excess reactant.

18. Which equation can be used to solve problems involving stoichiometry?
(1) $H_2(g) + I_2(g) \longrightarrow 2HI(g)$
(2) $H_2(g) + 2I_2(g) \longrightarrow 2H_2I(g)$
(3) $2H_2(g) + I_2(g) \longrightarrow 2HI(g)$
(4) $2H_2(g) + 2I_2(g) \longrightarrow 4HI(g)$

19. If a chemist calculates the maximum amount of product that might be obtained in a chemical reaction, he or she is calculating the
(1) theoretical yield.
(2) mole ratio.
(3) percentage yield.
(4) actual yield.

PART B–1

For each item, write on a separate piece of paper the number of the word, expression, or statement that best answers the item.

20. For the reaction shown below, what is minimum quantity of $O_2(g)$ that is required for the production of 6 mol of $CO_2(g)$?

$$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(g)$$

(1) 5 mol
(2) 6 mol
(3) 7.5 mol
(4) 10.5 mol

21. What is the minimum quantity of $O_2(g)$ required for the production of 18 mol of $CO_2(g)$ in the combustion reaction shown below?

$$C_6H_{12}(g) + 9O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(g)$$

(1) 3 mol
(2) 9 mol
(3) 24 mol
(4) 27 mol

Answer questions 22–24 based on the decomposition reaction shown below.

$$NH_4NO_3(s) \longrightarrow N_2O(g) + 2H_2O(g)$$

22. What is the molar mass of NH_4NO_3?
(1) 31 g
(2) 34 g
(3) 40 g
(4) 80 g

23. What is the minimum mass of $NH_4NO_3(s)$ that is required to produce 18 grams of $H_2O(g)$?
(1) 9 g
(2) 40 g
(3) 80 g
(4) 160 g

24. How many moles of $N_2O(g)$ are produced by the decomposition of 0.5 mol of $NH_4NO_3(s)$?
(1) 0.5 mol
(2) 1 mol
(3) 22 mol
(4) 44 mol

25. What are the limiting reactant and the excess reactant when firewood burns in a campfire?
(1) firewood is the limiting reactant and oxygen is the excess reactant
(2) oxygen is the limiting reactant and firewood is the excess reactant
(3) oxygen and firewood are limiting reactants
(4) neither oxygen and firewood are excess reactants

26. What is the theoretical yield of $NO_2(g)$, when 2 mol of $NO(g)$ reacts with 1 mol of $O_2(g)$ according to the equation below?

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

(1) 2 g
(2) 46 g
(3) 92 g
(4) 184 g

27. How is percent yield calculated for a reaction where the theoretical yield is 25 g and the actual yield is 20 g?

(1) $\frac{20}{25} \times 100$

(2) $\frac{25}{20} \times 100$

(3) $\frac{(25 - 20)}{25} \times 100$

(4) $\frac{(20 - 25)}{20} \times 100$

PART B–2

Answer the following items.

28. Examine the following reaction.

$$2NaN_3(s) \longrightarrow 2Na(s) + 3N_2(g)$$

a. How many moles of $N_2(g)$ can be formed from the decomposition of 2.7 mol of $NaN_3(s)$?

b. How many grams of Na are produced from the decomposition of 31.1 g of NaN_3(s)?

c. How many molecules of $N_2(g)$ are formed when 3.24×10^{23} atoms of $Na(s)$ are formed?

29. Examine the following reaction.

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

When 55 g of $NO(g)$ is mixed with 35 g of $O_2(g)$, 31 g of $NO_2(g)$ is produced.

a. Which is the limiting reactant?

b. What is the theoretical yield in grams of $NO_2(g)$?

c. What is the percentage yield of $NO_2(g)$?

30. Examine the following reaction.

$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

a. How many moles of oxygen are produced by the reaction of 4 mol of $KClO_3(s)$?

b. What volume of oxygen can be made from 5.00×10^{-2} mol of $KClO_3$? The density of $O_2(g)$ is 1.428 g/L.

c. How many grams of $KClO_3$ must react to produce 42.0 mL of O_2?

d. How many mL of $O_2(g)$ will form at STP from the reaction of 55.2 g of $KClO_3$?

e. What type of reaction is represented by the equation?

31. Sulfuric acid reacts with aluminum hydroxide by double replacement.

a. Write the balanced equation for the reaction that occurs between sulfuric acid and aluminum hydroxide.

b. The reaction goes to completion when 30.0 g of sulfuric acid is mixed with 25.0 g of aluminum hydroxide. Which is the limiting reactant?

c. Calculate the theoretical yield of aluminum sulfate in grams.

d. If 30.0 g of aluminum sulfate is produced in this reaction, what is the percent yield of aluminum sulfate?

32. Hydrogen peroxide, H_2O_2, decomposes as shown in the following equation.

$$2H_2O_2(l) \longrightarrow 2H_2O(l) + O_2(g)$$

How many liters of O_2 can be made from 342 g of H_2O_2 if the density of O_2 is 1.428 g/L?

PART C

Answer the following items.

33. In photosynthesis, plants use energy from the sun to produce glucose, $C_6H_{12}O_6$, and oxygen, O_2, from the reaction between carbon dioxide, CO_2, and water, H_2O.

a. Write the balanced chemical equation for photosynthesis.

b. What is the sum of all of the coefficients in the balanced equation?

c. During photosynthesis, how many moles of carbon dioxide are removed from the air in the production of 6 moles of oxygen?

34. The first step in the industrial manufacture of nitric acid is the catalytic oxidation of ammonia. The unbalanced equation is shown below.

$$NH_3(g) + O_2(g) \longrightarrow NO(g) + H_2O(g)$$

a. Balance this equation.

b. How many moles of $NO(g)$ are formed by the reaction of 8 mol of $O_2(g)$?

c. If the reaction uses 170 g of ammonia, how many moles of oxygen are needed?

d. How many grams of water are produced when 60 g of $NO(g)$ are produced?

35. Assume that 44.3 g of Na_2O are formed during the inflation of an air bag. How many liters of CO_2 are needed to completely react with the Na_2O? The density of CO_2 is 1.35 g/L. The reaction is shown below.

$$Na_2O(s) + 2CO_2(g) + H_2O(g) \longrightarrow 2NaHCO_3(s)$$

36. Assume that 59.5 L of $N_2(g)$ with a density of 0.92 g/L is needed to fill an air bag according to the following equation.

$$2NaN_3(s) \longrightarrow 2Na(s) + 3N_2(g)$$

a. What mass of NaN_3 is needed to form this volume of nitrogen?

b. How many liters of $N_2(g)$ are actually made from 65.7 g NaN_3 if the yield is 94%?

c. What mass of NaN_3 is actually needed to form 59.5 L of $N_2(g)$?

37. Write a balanced equation for the combustion of octane, C_8H_{18}, with oxygen to obtain carbon dioxide and water. What is the mole ratio of oxygen to octane?

38. What mass of oxygen is required to burn 688 g of octane, C_8H_{18}, completely?

39. How many liters of O_2 are needed for the complete combustion of 1.00 L of C_8H_{18}? The density of octane is 0.700 g/mL.

40. Nitrogen dioxide from a car's exhaust reacts with oxygen to form ozone as shown in the following equation.

$$NO_2(g) + O_2(g) \longrightarrow NO(g) + O_3(g)$$

What mass of ozone can be formed from 4.55g of NO_2? If only 4.58 g of O_3 is formed, what is the percentage yield of this reaction?

41. The following reaction can be used to remove CO_2 breathed out by astronauts in a spacecraft.

$$2LiOH(s) + CO_2(g) \longrightarrow Li_2CO_3(s) + H_2O(l)$$

How many grams of carbon dioxide can be removed by 5.5 mol LiOH?

Base your answers to questions 42–47 based on the following passage and on your knowledge of chemistry.

An Explosive Reaction

Explosives contain substances that when mixed together produce an extremely quick and highly exothermic reactions. An exothermic reaction is a reaction that releases energy as heat. The reactants in explosives should be relatively stable, but should decompose rapidly when the reaction is initiated. The reactants should have weak chemical bonds. The chemical reaction produces molecules that are much more stable than the reactants. Two common products of explosive reactions include nitrogen, N_2 and carbon dioxide, CO_2. Both N_2 and CO_2 contain multiple bonds, which are generally stronger and more stable than single bonds. Nitroglycerine, an explosive, decomposes exothermically into nitrogen, oxygen, carbon dioxide and water. The decomposition of nitroglycerine is shown below.

$$4C_3H_5N_3O_9(l) \longrightarrow 6N_2(g) + O_2(g) + 12CO_2(g) + 10H_2O(g) + \text{energy}$$

42. What are two characteristics of reactants in explosives?

43. In the decomposition of nitroglycerine, what is the mole ratio of reactant to products?

44. How many moles of nitroglycerine are needed to produce 15 mol of nitrogen gas?

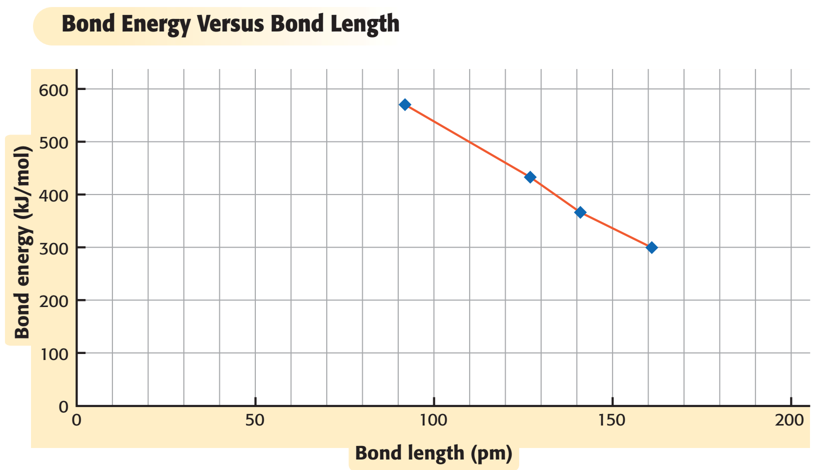
45. The molar mass of nitroglycerine is 227 g/mol. Calculate the mass of nitroglycerine that is needed to produce 90.0 g of water.

46. Calculate how many grams of oxygen are produced if 90.0 g of water forms.

47. Calculate the theoretical yield in grams of N_2 if 64 g of nitroglycerine are detonated.

FOCUS ON GRAPHING

Study the graph below, and answer the questions that follow.
For help in interpreting graphs, see Appendix B, "Study Skills for Chemistry."



48. Describe the relationship between bond length and bond energy.

49. Estimate the bond energy of a bond of length 100 pm.

50. If the trend of the graph continues, what bond length will have an energy of 200 kJ/mol?

51. The title of the graph does not provide much information about the contents of the graph. What additional information would be useful to better understand and use this graph?

TECHNOLOGY AND LEARNING

52. Graphing Calculator

Calculating Percentage Yield of a Chemical Reaction

The graphing calculator can run a program that calculates the percentage yield of a chemical reaction when you enter the actual yield and the theoretical yield. Using an example in which the actual yield is 38.8 g and the theoretical yield is 53.2 g, you will calculate the percentage yield. First, the program will carry out the calculation. Then you can use it to make other calculations.

Go to Appendix C. If you are using a TI-83 Plus, you can download the program YIELD and data and run the application as directed. If you are using another calculator, your teacher will provide you with keystrokes and data sets to use. After you have run the program, answer the questions.

Note: all answers are written with three significant figures.

a. What is the percentage yield when the actual yield is 27.3 g and the theoretical yield is 44.6 g?

b. What is the percentage yield when the actual yield is 5.40 g and the theoretical yield is 9.20 g?

c. What actual yield/theoretical yield pair produced the largest percentage yield?