12.3 Limiting Reagent and Percent Yield

Guide for Reading

Key Concepts

- How is the amount of product in a reaction affected by an insufficient quantity of any of the reactants?
- What does the percent yield of a reaction measure?

Vocabulary

limiting reagent
excess reagent
theoretical yield
actual yield
percent yield

Reading Strategy

Building Vocabulary After you have read the section, explain the differences among *theoretical yield, actual yield,* and *percent yield.*

Connecting to Your World

If a carpenter had two tabletops and seven table legs, he would have difficulty building more than one functional four-legged table. The first table would require four of the legs, leaving just three legs for the second table. In this case, the number of table legs is the limiting factor in the construction of four-legged tables. A similar concept applies in chemistry. The amount of product made in a chemical reaction may be limited by the amount of one or more of the reactants.

Limiting and Excess Reagents

Many cooks follow a recipe when making a new dish. They know that sufficient quantities of all the ingredients must be available. Suppose, for example, that you are preparing to make lasagna and you have more than enough meat, tomato sauce, ricotta cheese, eggs, mozzarella cheese, spinach, and seasoning on hand. However, you have only half a box of lasagna noodles. The amount of lasagna you can make will be limited by the quantity of noodles you have. Thus, the noodles are the limiting ingredient in this baking venture. Figure 12.9 illustrates another example of a limiting ingredient in the kitchen. A chemist often faces a similar situation. **In a chemical reaction, an insufficient quantity of any of the reactants will limit the amount of product that forms.**

Figure 12.9 The amount of product is determined by the quantity of the limiting reagent. In this example, the rolls are the limiting reagent. No matter how much of the other ingredients you have, with two rolls you can make only two sandwiches.

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Chemical Equations					
	$N_2(g)$	+	$3H_2(g)$	$\longrightarrow$	$2NH_3(g)$
"Microscopic recipe"	1 molecule N_2	+	3 molecules H_2	$\longrightarrow$	2 molecules NH_3
"Macroscopic recipe"	1 mol N_2	+	3 mol H_2	$\longrightarrow$	2 mol NH_3

Experimental Conditions

	Reactants		Products
Before reaction	 2 molecules N_2	3 molecules H_2	0 molecules NH_3
After reaction	 1 molecule N_2	0 molecules H_2	 2 molecules NH_3

Figure 12.10 The "recipe" calls for 3 molecules of H_2 for every 1 molecule of N_2. In this particular experiment, H_2 is the limiting reagent and N_2 is in excess. **Inferring** ***How would the amount of products formed change if you started with four molecules of N_2 and three molecules of H_2?***

As you know, a balanced chemical equation is a chemist's recipe. You can interpret the recipe on a microscopic scale (interacting particles) or on a macroscopic scale (interacting moles). The coefficients used to write the balanced equation give both the ratio of representative particles and the mole ratio. Recall the equation for the preparation of ammonia:

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

When one molecule (mole) of N_2 reacts with three molecules (moles) of H_2, two molecules (moles) of NH_3 are produced. What would happen if two molecules (moles) of N_2 reacted with three molecules (moles) of H_2? Would more than two molecules (moles) of NH_3 be formed? Figure 12.10 shows both the particle and the mole interpretations of this problem.

Before the reaction takes place, nitrogen and hydrogen are present in a 2:3 molecule (mole) ratio. The reaction takes place according to the balanced equation. One molecule (mole) of N_2 reacts with three molecules (moles) of H_2 to produce two molecules (moles) of NH_3. At this point, all the hydrogen has been used up, and the reaction stops. One molecule (mole) of unreacted nitrogen is left in addition to the two molecules (moles) of NH_3 that have been produced by the reaction.

In this reaction, only the hydrogen is completely used up. It is the **limiting reagent,** or the reagent that determines the amount of product that can be formed by a reaction. The reaction occurs only until the limiting reagent is used up. By contrast, the reactant that is not completely used up in a reaction is called the **excess reagent.** In this example, nitrogen is the excess reagent because some nitrogen will remain unreacted.

Sometimes in stoichiometric problems, the given quantities of reactants are expressed in units other than moles. In such cases, the first step in the solution is to convert each reactant to moles. Then the limiting reagent can be identified. The amount of product formed in a reaction can be determined from the given amount of limiting reagent.

Checkpoint ***How do limiting and excess reagents differ?***

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For: Links on Hydrogen
Visit: www.SciLinks.org
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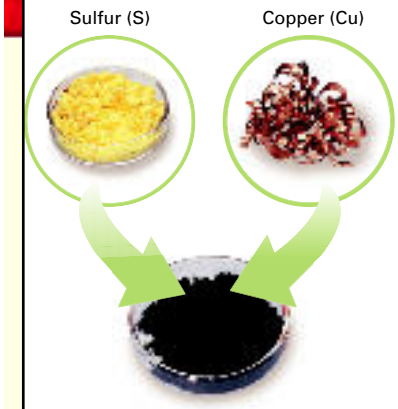
Interactive Textbook
Animation 13 Apply the limiting reagent concept to the production of iron from iron ore.
with **ChemASAP**



SAMPLE PROBLEM 12.7

Determining the Limiting Reagent in a Reaction

Copper reacts with sulfur to form copper(I) sulfide according to the following balanced equation.

$$2Cu(s) + S(s) \longrightarrow Cu_2S(s)$$

What is the limiting reagent when 80.0 g Cu reacts with 25.0 g S?

1 Analyze ***List the knowns and the unknown.***

Knowns
- mass of copper = 80.0 g Cu
- mass of sulfur = 25.0 g S

Unknown
- limiting reagent = ?

The number of moles of each reactant must first be found:

$$\text{g Cu} \longrightarrow \text{mol Cu}$$

$$\text{g S} \longrightarrow \text{mol S}$$

The balanced equation is used to calculate the number of moles of one reactant needed to react with the given amount of the other reactant:

$$\text{mol Cu} \longrightarrow \text{mol S}$$

The mole ratio relating mol S to mol Cu from the balanced chemical equation is 1 mol S/2 mol Cu.

2 Calculate ***Solve for the unknown.***

$$80.0\ \cancel{\text{g Cu}} \times \frac{1\ \text{mol Cu}}{63.5\ \cancel{\text{g Cu}}} = 1.26\ \text{mol Cu}$$

$$25.0\ \cancel{\text{g S}} \times \frac{1\ \text{mol S}}{32.1\ \cancel{\text{g S}}} = 0.779\ \text{mol S}$$

$$1.26\ \cancel{\text{mol Cu}} \times \frac{1\ \text{mol S}}{2\ \cancel{\text{mol Cu}}} = 0.630\ \text{mol S}$$

Given quantity — Mole ratio — Needed amount

Comparing the amount of sulfur needed (0.630 mol S) with the given amount (0.779 mol S) indicates that sulfur is in excess. Thus copper is the limiting reagent.

3 Evaluate ***Do the results make sense?***

Since the ratio of the given mol Cu to mol S was less than the ratio (2:1) from the balanced equation, copper should be the limiting reagent.

Math Handbook

For help with dimensional analysis, go to page R66.

Interactive Textbook

Problem-Solving 12.25 Solve Problem 25 with the help of an interactive guided tutorial.

with ChemASAP

Practice Problems

25. The equation for the complete combustion of ethene (C_2H_4) is

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(g)$$

If 2.70 mol C_2H_4 is reacted with 6.30 mol O_2, identify the limiting reagent.

26. Hydrogen gas can be produced by the reaction of magnesium metal with hydrochloric acid.

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

Identify the limiting reagent when 6.00 g HCl reacts with 5.00 g Mg.

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In Sample Problem 12.7, you may have noticed that even though the mass of copper used in the reaction is greater than the mass of sulfur, copper is the limiting reagent. The reactant that is present in the smaller amount by mass or volume is not necessarily the limiting reagent.

SAMPLE PROBLEM 12.8

Using a Limiting Reagent to Find the Quantity of a Product

What is the maximum number of grams of Cu_2S that can be formed when 80.0 g Cu reacts with 25.0 g S?

$$2Cu(s) + S(s) \longrightarrow Cu_2S(s)$$

1 Analyze ***List the knowns and the unknown.***

Knowns

- limiting reagent = 1.26 mol Cu (from Sample Problem 12.7)
- 1 mol Cu_2S = 159.1 g Cu_2S (molar mass)

Unknown

- mass copper(I) sulfide = ? g Cu_2S

The limiting reagent, which was determined in the previous sample problem, is used to calculate the maximum amount of Cu_2S formed:

$$\text{mol Cu} \longrightarrow \text{mol } Cu_2S \longrightarrow \text{g } Cu_2S$$

2 The equation yields the appropriate mole ratio: 1 mol Cu_2S/2 mol Cu.

Calculate ***Solve for the unknown.***

$$1.26\ \cancel{\text{mol Cu}} \times \frac{1\ \cancel{\text{mol } Cu_2S}}{2\ \cancel{\text{mol Cu}}} \times \frac{159.1\ \text{g } Cu_2S}{1\ \cancel{\text{mol } Cu_2S}} = 1.00 \times 10^2\ \text{g } Cu_2S$$

The given quantity of copper, 80.0 g, could have been used for this step instead of the moles of copper, which were calculated in Sample Problem 12.7.

3 Evaluate ***Do the results make sense?***

Copper is the limiting reagent in this reaction. The maximum number of grams of Cu_2S produced should be more than the amount of copper that initially reacted because copper is combining with sulfur. However, the mass of Cu_2S produced should be less than the total mass of the reactants (105.0 g) because sulfur was in excess.

Interactive Textbook

Problem-Solving 12.28 Solve Problem 28 with the help of an interactive guided tutorial.

with ChemASAP

Practice Problems

27. The equation below shows the incomplete combustion of ethene.

$$C_2H_4(g) + 2O_2(g) \longrightarrow 2CO(g) + 2H_2O(g)$$

If 2.70 mol C_2H_4 is reacted with 6.30 mol O_2,

a. identify the limiting reagent.

b. calculate the moles of water produced.

28. The heat from an acetylene torch is produced by burning acetylene (C_2H_2) in oxygen.

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

How many grams of water can be produced by the reaction of 2.40 mol C_2H_2 with 7.40 mol O_2?

Quick LAB

Limiting Reagents

Purpose

To illustrate the concept of a limiting reagent in a chemical reaction.

Materials

- graduated cylinder
- balance
- 3 250-mL Erlenmeyer flasks
- 3 rubber balloons
- 4.2 g magnesium ribbon
- 300 mL 1.0*M* hydrochloric acid

Procedure

1. Add 100 mL of the hydrochloric acid solution to each flask.
2. Weigh out 0.6 g, 1.2 g, and 2.4 g of magnesium ribbon, and place each sample into its own balloon.
3. Stretch the end of each balloon over the mouth of each flask. Do not allow the magnesium ribbon in the balloon to fall into the flask.
4. Magnesium reacts with hydrochloric acid to form hydrogen gas. When you mix the magnesium with the hydrochloric acid in the next step, you will generate a certain volume of hydrogen gas. How do you think the volume of hydrogen produced in each flask will compare?
5. Lift up on each balloon and shake the magnesium into each flask. Observe the volume of gas produced until the reaction in each flask is completed.

Analyze and Conclude

1. How did the volumes of hydrogen gas produced, as measured by the size of the balloons, compare? Did the results agree with your prediction?
2. Write a balanced equation for the reaction you observed.
3. The 100 mL of hydrochloric acid contained 0.10 mol HCl. Show by calculation why the balloon with 1.2 g Mg inflated to about twice the size of the balloon with 0.60 g Mg.
4. Show by calculation why the balloons with 1.2 g and 2.4 g Mg inflated to approximately the same volume. What was the limiting reagent when 2.4 g Mg was added to the acid?

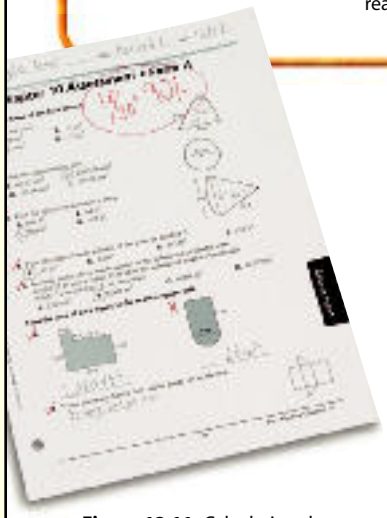

Figure 12.11 Calculating the ratio of the number of correct answers to the number of questions on the exam is a measure of how well the student performed on the exam.

Percent Yield

In theory, when a teacher gives an exam to the class, every student should get a grade of 100%. This generally does not occur, as shown in Figure 12.11. Instead, the performance of the class is usually spread over a range of grades. Your exam grade, expressed as a percentage, is a ratio of two items. The first item is the number of questions you answered correctly. The second is the total number of questions. The grade compares how well you performed with how well you could have performed if you had answered all the questions correctly. Chemists perform similar calculations in the laboratory when the product from a chemical reaction is less than expected, based on the balanced chemical equation.

When an equation is used to calculate the amount of product that will form during a reaction, the calculated value represents the theoretical yield. The **theoretical yield** is the maximum amount of product that could be formed from given amounts of reactants. In contrast, the amount of product that actually forms when the reaction is carried out in the laboratory is called the **actual yield.** The **percent yield** is the ratio of the actual yield to the theoretical yield expressed as a percent.

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

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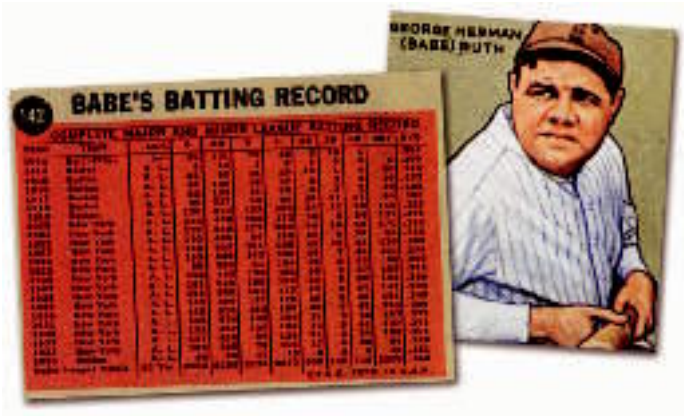


Figure 12.12 A batting average is actually a percent yield.

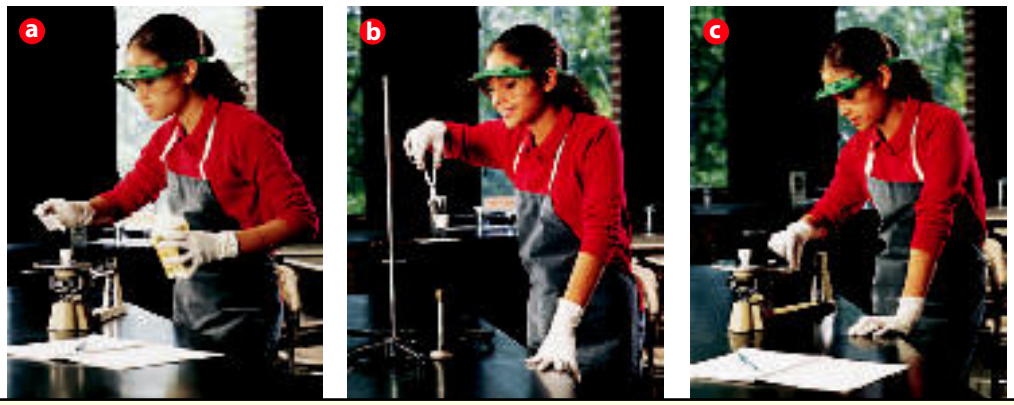
Because the actual yield of a chemical reaction is often less than the theoretical yield, the percent yield is often less than 100%. **The percent yield is a measure of the efficiency of a reaction carried out in the laboratory.** This is similar to an exam score measuring your efficiency of learning, or a batting average measuring your efficiency of hitting a baseball.

A percent yield should not normally be larger than 100%. Many factors cause percent yields to be less than 100%. Reactions do not always go to completion; when this occurs, less than the calculated amount of product is formed. Impure reactants and competing side reactions may cause unwanted products to form. Actual yield can also be lower than the theoretical yield due to a loss of product during filtration or in transferring between containers. Moreover, if reactants or products have not been carefully measured, a percent yield of 100% is unlikely.

An actual yield is an experimental value. Figure 12.13 shows a typical laboratory procedure for determining the actual yield of a product of a decomposition reaction. For reactions in which percent yields have been determined, you can calculate and therefore predict an actual yield if the reaction conditions remain the same.

Checkpoint ***What factors can cause the actual yield to be less than the theoretical yield?***

Figure 12.13 Sodium hydrogen carbonate ($NaHCO_3$) will decompose when heated. **a** The mass of $NaHCO_3$, the reactant, is measured. **b** The reactant is heated. **c** The mass of one of the products, sodium carbonate (Na_2CO_3), the actual yield, is measured. The percent yield is calculated once the actual yield is determined. **Predicting** ***What are the other products of this reaction?***



SAMPLE PROBLEM 12.9

Calculating the Theoretical Yield of a Reaction

Calcium carbonate, which is found in seashells, is decomposed by heating. The balanced equation for this reaction is:

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

What is the theoretical yield of CaO if 24.8 g $CaCO_3$ is heated?

1 Analyze ***List the knowns and the unknown.***

Knowns

- mass of calcium carbonate = 24.8 g $CaCO_3$
- 1 mol $CaCO_3$ = 100.1 g $CaCO_3$ (molar mass)
- 1 mol CaO = 56.1 g CaO (molar mass)

Unknown

- theoretical yield of calcium oxide = ? g CaO

Calculate the theoretical yield using the mass of the reactant:

$$\text{g } CaCO_3 \longrightarrow \text{mol } CaCO_3 \longrightarrow \text{mol CaO} \longrightarrow \text{g CaO}$$

The appropriate mole ratio is 1 mol CaO/1 mol $CaCO_3$.

2 Calculate ***Solve for the unknown.***

$$24.8 \text{ g } \cancel{CaCO_3} \times \frac{1 \text{ mol } \cancel{CaCO_3}}{100.1 \text{ g } \cancel{CaCO_3}} \times \frac{1 \text{ mol } \cancel{CaO}}{1 \text{ mol } \cancel{CaCO_3}} \times \frac{56.1 \text{ g CaO}}{1 \text{ mol } \cancel{CaO}}$$

$$= 13.9 \text{ g CaO}$$

3 Evaluate ***Does the result make sense?***

The mole ratio of CaO to $CaCO_3$ is 1:1. The ratio of their masses in the reaction should be the same as the ratio of their molar masses, which is slightly greater than 1:2. The result of the calculations shows that the mass of CaO is slightly greater than half the mass of $CaCO_3$.

Math Handbook

For help with dimensional analysis, go to page R66.

Practice Problems

29. When 84.8 g of iron(III) oxide reacts with an excess of carbon monoxide, iron is produced.

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g)$$

What is the theoretical yield of iron?

30. When 5.00 g of copper reacts with excess silver nitrate, silver metal and copper(II) nitrate are produced. What is the theoretical yield of silver in this reaction?

Interactive Textbook

Problem-Solving 12.29 Solve Problem 29 with the help of an interactive guided tutorial.

with ChemASAP

Recall that the percent yield is calculated by multiplying the ratio of the actual yield to theoretical yield by 100%. Therefore, you must have values of both the theoretical yield and the actual yield to calculate the percent yield.

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SAMPLE PROBLEM 12.10

Calculating the Percent Yield of a Reaction

What is the percent yield if 13.1 g CaO is actually produced when 24.8 g $CaCO_3$ is heated?

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

1 Analyze ***List the knowns and the unknown.***

Knowns

- actual yield = 13.1 g CaO
- theoretical yield = 13.9 g CaO (from Sample Problem 12.9)
- $\text{percent yield} = \dfrac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$

Unknown

- percent yield = ? %

2 Calculate ***Solve for the unknown.***

$$\text{percent yield} = \frac{13.1\ \cancel{\text{g CaO}}}{13.9\ \cancel{\text{g CaO}}} \times 100\% = 94.2\%$$

3 Evaluate ***Does the result make sense?***

In this example, the actual yield is slightly less than theoretical yield. Therefore, the percent yield should be slightly less than 100%. The answer should have three significant figures.

Practice Problems

31. If 50.0 g of silicon dioxide is heated with an excess of carbon, 27.9 g of silicon carbide is produced.

$$SiO_2(s) + 3C(s) \longrightarrow SiC(s) + 2CO(g)$$

What is the percent yield of this reaction?

32. If 15.0 g of nitrogen reacts with 15.0 g of hydrogen, 10.5 g of ammonia is produced. What is the percent yield of this reaction?

Math Handbook

For help with percents, go to page R72.

Interactive Textbook

Problem-Solving 12.31 Solve Problem 31 with the help of an interactive guided tutorial.

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12.3 Section Assessment

33. **Key Concept** In a chemical reaction, how does an insufficient quantity of a reactant affect the amount of product formed?

34. **Key Concept** How can you gauge the efficiency of a reaction carried out in the laboratory?

35. What is the percent yield if 4.65 g of copper is produced when 1.87 g of aluminum reacts with an excess of copper(II) sulfate?

$$2Al(s) + 3CuSO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3Cu(s)$$

Elements Handbook

Haber Process Read about ammonia on page R26. Examine the flow chart summarizing the Haber process. What experimental data would you need to determine the percent yield of the Haber process?

Interactive Textbook

Assessment 12.3 Test yourself on the concepts in Section 12.3.

with ChemASAP

Technology & Society

Just the Right Volume of Gas

In a front-end collision, proper inflation of an air bag may save your life. Engineers use stoichiometry to determine the exact quantity of each reactant in the air bag's inflation system. **Interpreting Diagrams** ***What is the source of the gas that fills an air bag?***

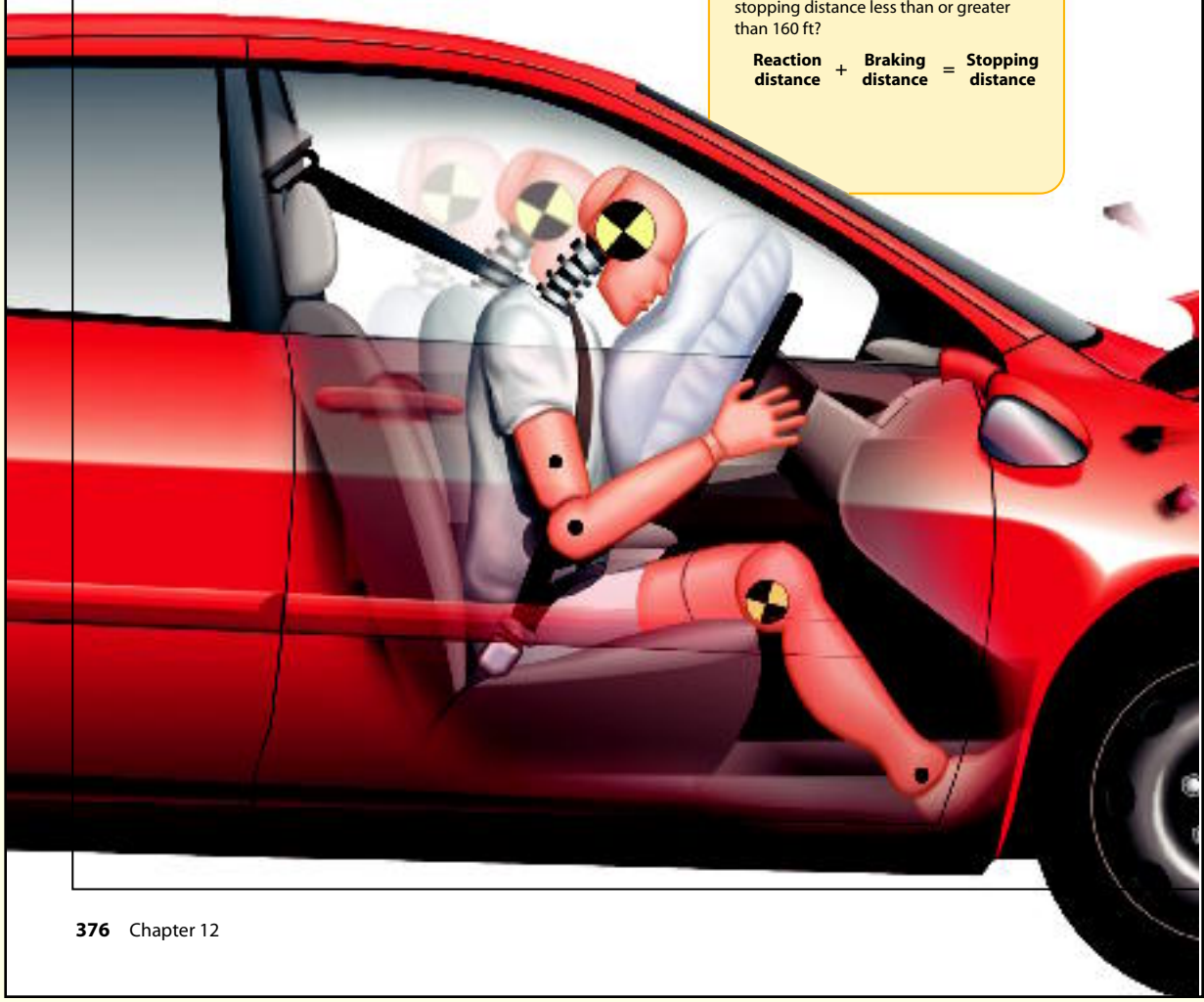
Car Facts

17.1 million cars and light trucks were sold in 2002 in the United States.

Monaco has the highest number of vehicles in relation to its road network. In 1996 (most recent figures), it had 480 vehicles for each kilometer of road. If they were required to park behind one another on the streets, half would have nowhere to park!

At night, headlights illuminate 160 ft in front of your car. If you are driving 40 mi/h at night, your reaction distance is 88 ft and your braking distance is 101 ft. Is your stopping distance less than or greater than 160 ft?

Reaction distance + **Braking distance** = **Stopping distance**



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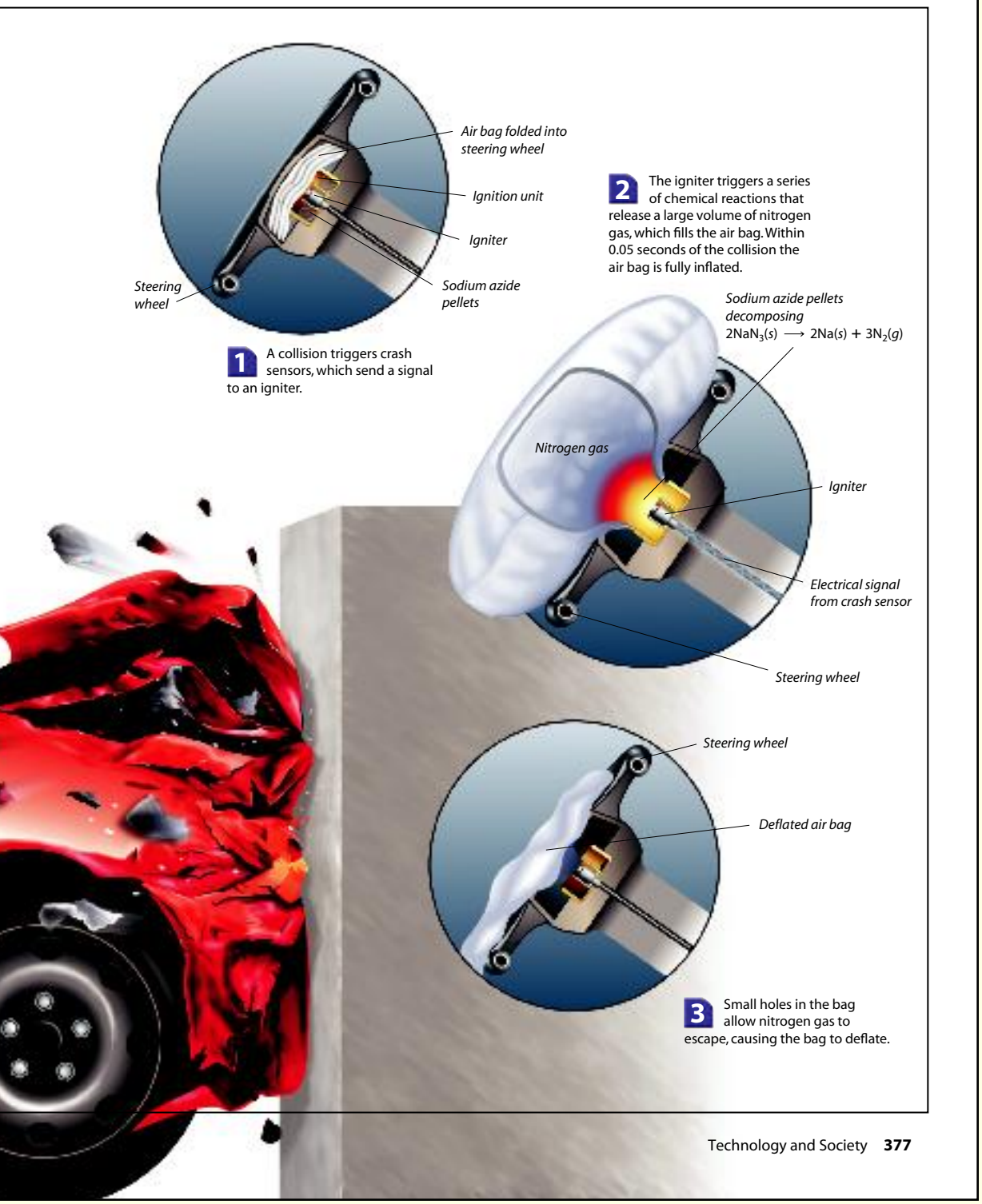

1 A collision triggers crash sensors, which send a signal to an igniter.

2 The igniter triggers a series of chemical reactions that release a large volume of nitrogen gas, which fills the air bag. Within 0.05 seconds of the collision the air bag is fully inflated.

3 Small holes in the bag allow nitrogen gas to escape, causing the bag to deflate.