Mass Relationships in Chemical Equations

7.2

There are many dangers aboard a spacecraft. One of the more mundane is the possibility of too much carbon dioxide in the air. Carbon dioxide can be toxic if it is allowed to build up in a confined space. Too much carbon dioxide interferes with the oxygen delivery system in the body. This would make the astronauts sleepy, cause their breathing rate to increase, and lower their blood pH. The source of the carbon dioxide in a spacecraft is the air exhaled by the astronauts. Fortunately for the astronauts, air-scrubbing technology aboard spacecraft removes much of the carbon dioxide before it can accumulate to dangerous levels (**Figure 1**).

WEB **LINK**

In 1970, when space exploration was at its height, *Apollo 13* experienced an equipment failure that could have led to a fatal buildup of carbon dioxide in the spacecraft. To find out how the astronauts solved this problem,

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Figure 1 An astronaut changes a lithium hydroxide canister used to scrub carbon dioxide from the interior air during space missions. Each canister lasts about 11 hours for a seven-person crew.

This technology involves passing air through canisters of lithium hydroxide, LiOH. This compound reacts with carbon dioxide in the air to form lithium carbonate and water:

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

More common hydroxides like calcium hydroxide also undergo this reaction (**Figure 2**). However, lithium hydroxide is preferred for use in space because the mass of the lithium ion is a fraction of the mass of the calcium ion. Keeping mass, and therefore weight, to a minimum is a critical consideration when deciding what to bring onto a spacecraft. Prior to each mission, chemists must determine the quantity of lithium hydroxide required—just enough to scrub exhaled carbon dioxide, plus a little extra in case of flight delays. Any more would be "excess baggage." The required amount is based on the predicted daily output of carbon dioxide by the astronauts. As you will see in this section, the balanced chemical equation for this reaction plays a key role in determining the amount of lithium hydroxide required.

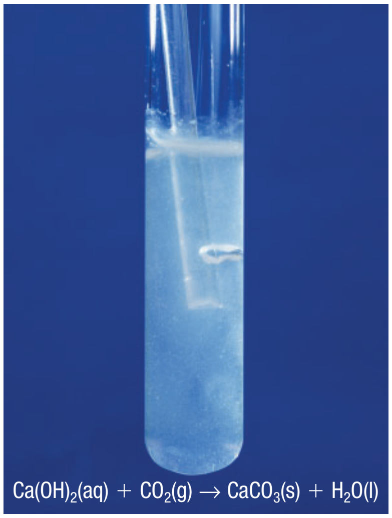

Figure 2 The addition of carbon dioxide makes limewater become cloudy due to the formation of a calcium carbonate precipitate. This reaction is commonly used in high school investigations to test for carbon dioxide.

Masses and Chemical Equations

SKILLS HANDBOOK A6

The coefficients of a balanced chemical equation give the mole ratio of one chemical to another. These values give the relative amount of one chemical required to react with another. They can also be used to predict the amount of product expected.

In Chapter 6, you learned that chemists use mass as a convenient way to determine the number of entities in a sample of a pure substance. The amount present is determined by dividing the mass of the substance by its molar mass ($n = \frac{m}{M}$). The study of the relationships between the amounts of reactants and products in chemical reactions is called **stoichiometry**.

stoichiometry the study of the mass and amount relationships between reactants and products in a chemical reaction

Tutorial 1 Using Stoichiometry to Find Mass

A typical astronaut exhales 8.8×10^2 g of carbon dioxide daily. Determining the mass of lithium hydroxide required to react with this mass of CO_2 begins with an analysis of the chemical equation:

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

Given: $m_{CO_2} = 8.8 \times 10^2$ g

Required: mass of lithium hydroxide, m_{LiOH}

Solution:

Step 1. Write a balanced equation for the reaction, listing the given value(s), required value(s), and molar masses below the substances being considered in the problem. Use the symbol of the required value since its value is unknown.

$CO_2(g)$ +	$2\ LiOH(s) \rightarrow Li_2CO_3(aq) + 2\ H_2O(l)$
8.8×10^2 g	m_{LiOH}
44.01 g/mol	23.95 g/mol

Step 2. Convert mass of given substance(s) to amount(s).

$$n_{CO_2} = 8.80 \times 10^2\ \cancel{g} \times \frac{1\ \text{mol}}{44.01\ \cancel{g}}$$

$$n_{CO_2} = 19.995\ \text{mol}\ [2\ \text{extra digits carried}]$$

Step 3. Convert amount of given substance to amount of required substance. To do this, multiply the amount of the given substance by a suitable conversion factor derived from the mole ratio in the balanced equation. The mole ratio between CO_2 and LiOH may be expressed as

$$\frac{2\ \text{mol}_{LiOH}}{1\ \text{mol}_{CO_2}} \text{ or } \frac{1\ \text{mol}_{CO_2}}{2\ \text{mol}_{LiOH}}$$

Since we want to convert amount of carbon dioxide to amount of lithium hydroxide, use the factor $\frac{2\ \text{mol}_{LiOH}}{1\ \text{mol}_{CO_2}}$ as follows:

$$n_{LiOH} = 19.995\ \cancel{\text{mol}_{CO_2}} \times \frac{2\ \text{mol}_{LiOH}}{1\ \cancel{\text{mol}_{CO_2}}}$$

$$n_{LiOH} = 39.991\ \text{mol}\ [2\ \text{extra digits carried}]$$

Step 4. Convert amount of required substance to mass of required substance.

$$m_{LiOH} = (39.991\ \cancel{\text{mol}})\left(\frac{23.95\ \text{g}}{1\ \cancel{\text{mol}}}\right)$$

$$m_{LiOH} = 958\ \text{g}$$

Statement: 958 g of lithium hydroxide is required to react with 8.8×10^2 g carbon dioxide.

Figure 3 summarizes the steps used in this problem.

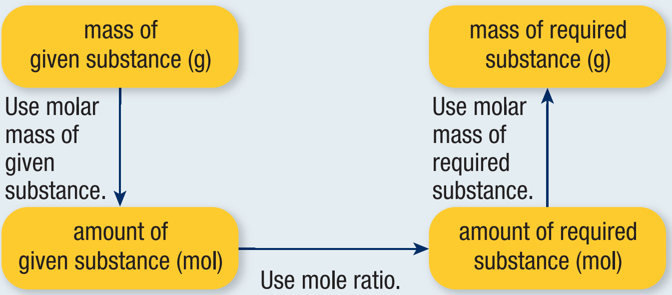

Figure 3 Solving a stoichiometry problem

The strategy used in the lithium hydroxide example can be generalized to solve most stoichiometric problems. As Figure 3 shows, key steps involved in solving stoichiometry problems involving masses are as follows:

1. Convert the given mass to an amount.
2. Apply the mole ratio of the two substances given in the question to predict the required amount of the other substance.
3. Convert the predicted amount to mass.

After completing your calculation, check that the number of significant digits in the final answer agrees with the least number of significant digits in the calculation. Also check that your final answer is numerically correct and that it has the correct unit.

Sample Problem 1: Calculating a Required Mass

An automobile airbag is inflated with nitrogen produced from the decomposition of sodium azide, NaN_3 (**Figure 4**):

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

The mass of nitrogen in a fully inflated airbag is 87.5 g. What mass of sodium azide is required to produce this mass of nitrogen?

Given: $m_{N_2} = 87.5$ g

Required: mass of sodium azide, m_{NaN_3}

Solution:

Step 1. Write a balanced equation listing given value(s), required value(s), and the corresponding molar masses.

$2\ NaN_3(s) \rightarrow$	$2\ Na(s) + 3\ N_2(g)$
m_{NaN_3}	87.5 g
65.02 g/mol	28.02 g/mol

Step 2. Convert mass of given substance to amount of given substance.

$$n_{N_2} = 87.5\ \cancel{g} \times \frac{1\ mol_{N_2}}{28.02\ \cancel{g}}$$

$$n_{N_2} = 3.1228\ [\text{2 extra digits carried}]$$

Step 3. Convert amount of given substance to amount of required substance.

$$n_{NaN_3} = 3.1228\ \cancel{mol_{N_2}} \times \frac{2\ mol_{NaN_3}}{3\ \cancel{mol_{N_2}}}$$

$$n_{NaN_3} = 2.0819\ mol$$

Step 4. Convert amount of required substance to mass of required substance.

$$m_{NaN_3} = (2.0819\ \cancel{mol_{NaN_3}})\left(\frac{65.02\ g}{1\ \cancel{mol_{NaN_3}}}\right)$$

$$m_{NaN_3} = 135\ g$$

Statement: The mass of sodium azide required to inflate the airbag is 135 g.

Figure 4 An automobile airbag is designed to inflate in a fraction of a second.

WEB LINK

It is very important that the correct mass of sodium azide is present in an airbag to ensure that it expands to the optimal size. The airbag must then quickly deflate to cushion the motorist's impact. To learn more about the design and function of airbags,

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Practice

SKILLS HANDBOOK A6

1. A typical antacid tablet contains 0.50 g of calcium carbonate. The chemical equation for the neutralization of hydrochloric acid with calcium carbonate is

$$CaCO_3(s) + 2\ HCl(aq) \rightarrow CO_2(g) + H_2O(l) + CaCl_2(aq)$$ T/I

(a) What mass of hydrochloric acid will this mass of calcium carbonate neutralize? [ans: 0.36 g]

(b) Predict what mass of calcium chloride will be produced. [ans: 0.55 g]

2. The chemical equation for the complete combustion of propane, $C_3H_8(g)$, is

$$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$$ T/I

(a) What mass of oxygen is required to burn 8.8 kg of propane? [ans: 32 kg]

(b) Predict what mass of carbon dioxide will be produced. [ans: 26 kg]

3. Freshly cut sodium undergoes a synthesis reaction with oxygen in the air. T/I

(a) Write the balanced chemical equation for this reaction.

(b) What mass of oxygen is required to react completely with 3.45 g of sodium? [ans: 1.20 g]

Stoichiometric Amounts

In the lithium hydroxide example at the beginning of Tutorial 1, 40.0 mol of lithium hydroxide is required to absorb 20.0 mol of carbon dioxide. Note that these amounts are multiples of the amounts given in the balanced chemical equation for the reaction. The quantities 40.0 mol LiOH and 20.0 mol CO_2 are stoichiometric amounts. A **stoichiometric amount** is the predicted amount of a reactant, relative to another reactant, that will react according to the balanced chemical equation. The amounts of reactants consumed are in the same proportion as the coefficients in the balanced chemical equation. When stoichiometric amounts of reactants are available for a chemical reaction, no reactants should remain when the reaction is complete. (This assumes that the reaction proceeds to completion.)

stoichiometric amount an amount of reactants that is in the same proportion as the reactant coefficients in the balanced chemical equation

It is important to be aware of stoichiometric amounts when designing any process involving chemical reactions. For example, in Unit 2 you learned that toxic emissions result from the combustion of fossil fuels. Calcium carbonate can be used to scrub sulfur dioxide from these emissions. The chemical reaction occurring in this type of scrubber is

$$SO_2(g) + CaCO_3(s) \rightarrow CaSO_3(s) + CO_2(g)$$

Note that sulfur dioxide and calcium carbonate are in a 1:1 ratio. As a result, sulfur dioxide and calcium carbonate are present in stoichiometric amounts in any mixture containing an equal amount of each compound. (Remember that we are considering equal amounts here, not equal masses.) If less than the stoichiometric amount of calcium carbonate is used, some unused sulfur dioxide escapes into the atmosphere where it can form acid precipitation. In practice, an excess of calcium carbonate is always present to ensure all the sulfur dioxide is captured.

As you will see in Section 7.3, when non-stoichiometric amounts are combined, one reactant remains when the reaction is complete.

Investigation 7.2.1

What Is Baking Soda Doing in Your Cake? (p. 340)
In this investigation you will decompose baking soda, determine the mass of solid product, and decide which of four possible reactions actually occurred. You will test the gases produced to confirm your conclusion.

7.2 Summary

- Stoichiometry makes use of the relationships between mass and amount of the reactants and products in a chemical reaction.
- The amount of products in a chemical reaction can be predicted from the amount of reactants. The masses of products cannot be predicted directly from the masses of reactants.
- Stoichiometry problems involving masses can be solved by converting masses to amounts and using mole ratios (Figure 3).
- Reactants are present in stoichiometric amounts if they are in the same ratio as given by the balanced chemical equation. Theoretically, stoichiometric amounts of reactants are completely used up in a chemical reaction.

7.2 Questions

1. Why is a balanced chemical equation required before solving a stoichiometry problem? K/U
2. A student adds vinegar (a solution of ethanoic acid) to baking soda (sodium hydrogen carbonate). After the bubbling subsides, a clear colourless solution remains. The student then adds more vinegar, producing more bubbles. Were ethanoic acid and sodium hydrogen carbonate initially present in stoichiometric amounts? Explain. T/I
3. Aluminum undergoes a synthesis reaction with oxygen to form aluminum oxide:

   $4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$

   Which of the following are stoichiometric amounts of the two reactants? K/U T/I

Quantity of Al(s)	**Quantity of O_2(g)**
2.0 mol	1.5 mol
0.60 mol	0.45 mol
0.13 mol	0.12 mol
108 g	96 g
4.0 g	3.0 g

4. The lithium ion battery (**Figure 5**) is an ideal power source for cellphones and digital cameras because its energy output is high for a relatively small mass. A reaction inside the battery generates electrical energy:

   $Li(s) + CoO_2(s) \rightarrow LiCoO_2(s)$ T/I

Figure 5 A lithium ion battery

   (a) When the battery is almost completely drained of power, almost 80 % of the initial amount of lithium metal is converted to $LiCoO_2(s)$. If the initial amounts of the reactants are equal, what percentage of $CoO_2(s)$ remains? Explain your answer.
   (b) When the battery is placed in a recharger, electrical energy is used to reverse the given reaction. How do the relative amounts of lithium and cobalt oxide change as the battery is recharged? Why?

5. Calcium oxide (lime), CaO(s), is a key ingredient in cement. This compound can be made by decomposing calcium hydroxide:

   $Ca(OH)_2(s) \rightarrow CaO(s) + H_2O(g)$ T/I

   (a) Predict what mass of calcium oxide will be produced when 1.00 g of calcium hydroxide is decomposed.
   (b) Predict the mass of water expected.
   (c) Is stoichiometry required to solve (b)? Explain your answer.
   (d) What law of chemical reactions do your answers to (a) and (b) illustrate? Explain.
   (e) A student incorrectly predicted that 1.00 g of calcium hydroxide produces 1.00 g of calcium oxide and 1.00 g of water. Why is this prediction incorrect?
6. Bubbling chlorine gas through a solution of potassium iodide results in the formation of elemental iodine:

   $Cl_2(g) + 2\ KI(aq) \rightarrow I_2(s) + 2\ KCl(aq)$ K/U T/I

   (a) Classify the type of reaction involved.
   (b) What mass of iodine can be produced from 1.5 kg of chlorine?
7. Iron undergoes a synthesis reaction to form iron(III) oxide when it is heated in air. K/U T/I C
   (a) Write a balanced chemical equation for this reaction.
   (b) What amount of iron reacts with 15 mol of oxygen? What amount of iron(III) oxide is produced?
   (c) If 4.60 g of iron(III) oxide is produced, what mass of iron reacted?
   (d) What is the total mass of iron and oxygen that reacted?
   (e) Is stoichiometry required to solve (d)? Explain your answer.
8. Ethanol, $C_2H_6O(l)$, is an important additive in gasoline. T/I A
   (a) Determine the mass of carbon dioxide produced during the complete combustion of 1.15 g of ethanol.
   (b) What might be the environmental consequences if insufficient oxygen were present for the combustion reaction?
9. The addition of heat causes ammonium nitrate, NH_4NO_3, to form dinitrogen monoxide (laughing gas), N_2O, and water. K/U T/I C
   (a) Write a balanced chemical equation for this reaction.
   (b) Classify this reaction.
   (c) Predict what mass of N_2O is produced when 1.00 g of ammonium nitrate reacts.
10. Nitric acid, $HNO_3(aq)$, can be manufactured from ammonia using a series of three chemical reactions called the Ostwald process. The reactions involved are

    $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$

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

    $3\ NO_2(g) + H_2O(g) \rightarrow 2\ HNO_3(aq) + NO(g)$

    Determine the mass of nitric acid produced if 425 kg of ammonia reacts. Assume that plenty of oxygen is available. T/I