12.5

# Reactions of Gases and Gas Stoichiometry

**Figure 1** The combustion reaction in a motorbike's engine produces a large volume of gases: the exhaust.

The roar of a motorcycle engine turns heads as it picks up speed and races past you. (**Figure 1**). What causes that sound and results in that impressive speed? A gas reaction!

Gases are involved in chemical reactions all around us, from medical applications to the gases used to heat our homes and power our vehicles. Gases interact chemically with each other and with other forms of matter. These interactions can be represented as balanced chemical reactions. The chemical reaction in the motorcycle engine involved gasoline vapour and oxygen, and produced mostly carbon dioxide and water. In this section we will examine chemical reactions involving gases and the volumes of those gases.

## Gas Stoichiometry

You learned a number of key concepts about stoichiometric relationships in chemical reactions in Unit 3. You know that chemical reactions involve mole ratios of chemicals. When working with chemical reactions where some of the reactants and/or products are gases, it is more likely that you will be measuring volumes rather than masses. We therefore use slightly different techniques to solve these stoichiometry problems.

Recall Gay-Lussac's law of combining volumes (Section 12.1). This law explains that, when gases react, the volumes of the reactants and products react in whole-number ratios if the temperatures and pressures are constant. This is useful for determining unknown volumes when working on gas stoichiometry problems.

### Tutorial 1 Volume-to-Volume Stoichiometry

When solving stoichiometry problems you often know the volumes of gaseous reactants and/or products and are asked to determine other volumes in the reaction. (This assumes that the temperature and pressure remain constant throughout the problem.) For this type of problem you will need to start by writing the balanced chemical equation. You will find it helpful to list any known quantities, with the appropriate units, underneath the equation. Make sure that similar variables have identical units. You can then use a conversion factor derived from the mole ratio of the two gases to calculate the required volume.

WEB **LINK**

To learn more about how catalytic converters convert carbon monoxide into carbon dioxide,

GO TO NELSON SCIENCE

#### Sample Problem 1: Using a Volume Ratio to Determine a Volume

A catalytic converter in the exhaust system of a car uses oxygen (from the air) and a catalyst to convert carbon monoxide to carbon dioxide (**Figure 2**). If the temperature and pressure remain the same, what volume of oxygen is required to react with 65.0 L of carbon monoxide produced during a road trip?

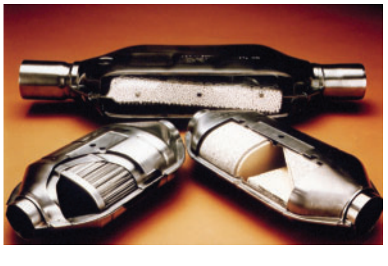

**Figure 2** A catalytic converter is part of the exhaust system in an automobile.

**Given:** volume of carbon monoxide, $V_{CO} = 65.0$ L

**Required:** volume of oxygen, $V_{O_2}$

**Solution:**

**Step 1.** Write the balanced chemical equation for the reaction, listing the given value(s) and required values below.

$2\ CO(g) + O_2(g) \rightarrow 2\ CO_2(g)$

65.0 L $\quad V_{O_2}$

**Step 2.** Convert the volume of the given substance, $V_{CO}$, to volume of the required substance, $V_{O_2}$. To do this, multiply by a conversion factor derived from the mole ratio of the given substance to the required substance. Since we want an answer involving oxygen, the conversion factor is $\frac{1\ mol_{O_2}}{2\ mol_{CO}}$.

$$V_{O_2} = V_{CO} \times \frac{1\ \text{mol}_{O_2}}{2\ \text{mol}_{CO}}$$

$$= 65.0\ \text{L} \times \frac{1\ \cancel{\text{mol}_{O_2}}}{2\ \cancel{\text{mol}_{CO}}}$$

$$V_{O_2} = 32.5\ \text{L}$$

**Statement:** The volume of oxygen required is 32.5 L.

## Practice

1. What volume of oxygen will be required for the complete combustion of 54 L of hydrogen if both gases are measured at the same temperature and pressure?

   $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$ T/I [ans: 27 L]
2. What volume of hydrogen gas, $H_2(g)$, is required to react with nitrogen gas, $N_2(g)$, to produce 34.5 L of ammonia, $NH_3(g)$? Assume that the temperature and pressure remain constant throughout the reaction. T/I [ans: 51.8 L]
3. Gunpowder is a mixture of potassium nitrate (commonly known as saltpetre), $KNO_3(s)$, charcoal, C(s), and sulfur, $S_8(s)$. When heated or struck by a sharp blow, the potassium nitrate decomposes to produce oxygen, which reacts rapidly with the charcoal and sulfur. The decomposition of saltpetre is shown by the following balanced chemical equation:

   $4\ KNO_3(s) \rightarrow 2\ K_2O(s) + 2\ N_2(g) + 5\ O_2(g)$

   Both gases are measured at the same temperature and pressure. What volume of oxygen is produced along with 15.0 L of nitrogen? T/I [ans: 37.5 L]

More commonly, temperature and pressure vary during a chemical reaction. In this case, we must use the ideal gas law equation ($PV = nRT$) to calculate the volume of a gas given the amount determined by stoichiometric techniques.

Airbags are a dramatic example of how a gas can save human lives (**Figure 3**). During a collision, sensors activate a chemical reaction that produces a volume of nitrogen gas that immediately inflates the airbag. After cushioning the driver, the airbag quickly deflates as nitrogen molecules escape through the permeable cover.

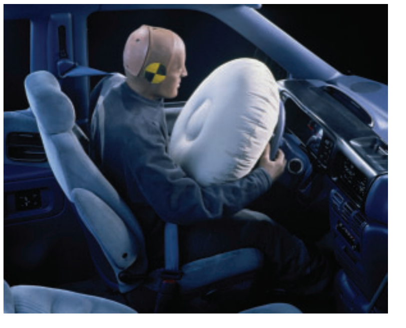

**Figure 3** A safety airbag uses gas stoichiometry to save lives.

The nitrogen gas is produced from a series of chemical reactions. The gas generator contains an electrical igniter and a precise mixture of three compounds: sodium azide, $NaN_3$; iron(III) oxide, $Fe_2O_3$; and silicon dioxide, $SiO_2$. This mixture is placed in a porous folded pouch. When ignited, the sodium azide decomposes very quickly to produce sodium metal and nitrogen gas:

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

Almost instantly, the sodium metal reacts with the iron(III) oxide to produce solid sodium oxide and iron:

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

Finally, the thermal energy released by these reactions melts the solid products and the silicon dioxide to form small pieces of a safe, unreactive solid similar to glass.

The nitrogen gas in the airbag must inflate to approximately 67.0 L at a certain pressure in order for the airbag to be safe and effective. Automobile designers and engineers need to determine what quantities of the various reactants are required to achieve this inflation volume. Airbags are designed to slow the forward motion of the average adult male. Children are generally smaller, so they are at risk of serious injury if they are in the front seat when an airbag inflates.

The same stoichiometric principles that you developed in Unit 3, combined with an understanding of the ideal gas law equation, will allow you to solve problems similar to the airbag situation.

**WEB LINK**

To learn more about how air bags work,

GO TO NELSON SCIENCE

## Tutorial 2 Mass or Volume-to-Mass or Volume Stoichiometry

In many stoichiometry problems you are given the mass, amount, or volume of gaseous reactants and/or products and asked to determine a corresponding quantity. Again, you will need the balanced chemical equation to determine the mole ratios. You may have to use molar mass, molar volume, the ideal gas law, or other appropriate concepts or equations as part of the solution. Make sure that your answer makes sense and is stated in the required units.

### **Sample Problem 1:** Determining the Volume of a Gaseous Product

What volume of carbon dioxide is produced when 6.40 g of methane gas, $CH_4(g)$, reacts with excess oxygen? All gases are at 35.0 °C and 100.0 kPa.

**Given:** mass of methane, $m_{CH_4} = 6.40$ g

molar mass of methane, $M_{CH_4} = 16.05$ g/mol

$t = 35.0$ °C and

$P = 100.0$ kPa.

**Required:** volume of carbon dioxide, $V_{CO_2}$

**Solution:**

**Step 1.** Convert the temperature value(s) to kelvins.

$$T = 35.0 + 273$$

$$T = 308 \text{ K}$$

**Step 2.** Write the balanced chemical equation, listing the given and required quantities, with the appropriate units, underneath.

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$$

6.40 g $\qquad\qquad V_{CO_2}$

**Step 3.** Convert the mass of the given substance, $m_{CH_4}$, into an amount, $n_{CH_4}$, using a conversion factor derived from the molar mass of methane, $CH_4$.

$$n_{CH_4} = 6.40\ \cancel{\text{g}} \times \frac{1 \text{ mol}}{16.05\ \cancel{\text{g}}}$$

$$n_{CH_4} = 0.3988 \text{ mol (extra digits carried)}$$

**Step 4.** Determine the amount of carbon dioxide produced from the amount of methane used using the appropriate mole ratio derived from the balanced chemical equation.

$$n_{CO_2} = 0.3988\ \cancel{\text{mol}_{CH_4}} \times \frac{1 \text{ mol}_{CO_2}}{1\ \cancel{\text{mol}_{CH_4}}}$$

$$n_{CO_2} = 0.3988 \text{ mol}_{CO_2}$$

**Step 5.** Determine the volume of the required substance, $V_{CO_2}$, using the ideal gas law equation, $PV = nRT$. Remember that $R = 8.314 \text{ kPa} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$.

$$PV = nRT$$

$$PV_{CO_2} = n_{CO_2}RT$$

$$V_{CO_2} = \frac{n_{CO_2}RT}{P}$$

$$= \frac{(0.3988\ \cancel{\text{mol}})\ (8.314\ \cancel{\text{kPa}} \cdot \text{L} \cdot \cancel{\text{mol}^{-1}} \cdot \cancel{\text{K}^{-1}})\ (308\ \cancel{\text{K}})}{100.0\ \cancel{\text{kPa}}}$$

$$V_{CO_2} = 10.2 \text{ L}$$

**Statement:** When 6.40 g of methane reacts with excess oxygen, 10.2 L of carbon dioxide is produced (under the given conditions of temperature and pressure).

### **Sample Problem 2:** Determining the Mass of a Gaseous Reactant

What mass of sodium azide is required to produce the 67.0 L of nitrogen gas that is required to fill a safety airbag in an automobile? Assume that the gas is produced at a temperature of 32 °C and a pressure of 105 kPa.

**Given:** volume of nitrogen gas, $V_{N_2} = 67.0$ L

temperature, $t = 32$ °C

pressure, $P = 105$ kPa

**Required:** mass of sodium azide, $m_{NaN_3}$

**Solution:**

**Step 1.** Convert the temperature value(s) to kelvins.

$$T = 32 + 273$$
$$T = 305 \text{ K}$$

**Step 2.** Write the balanced chemical equation, listing the given and required quantities, with the appropriate units, underneath.

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

$m_{NaN_3}$ 67.0 L

**Step 3.** Determine the amount of the given substance, $n_{N_2}$, produced in the reaction using the ideal gas law equation, $PV = nRT$.

$$PV = nRT$$
$$PV_{N_2} = n_{N_2}RT$$
$$n_{N_2} = \frac{PV_{N_2}}{RT}$$
$$= \frac{(105\ \cancel{\text{kPa}})(67.0\ \cancel{\text{L}})}{(8.314\ \cancel{\text{kPa}} \cdot \cancel{\text{L}} \cdot \text{mol}^{-1} \cdot \cancel{\text{K}^{-1}})(305\ \cancel{\text{K}})}$$
$$n_{N_2} = 2.774 \text{ mol (extra digits carried)}$$

**Step 4.** Determine the amount of sodium azide used to produce the amount of nitrogen gas determined in Step 3, using the appropriate mole ratio derived from the balanced chemical equation:

$$n_{NaN_3} = 2.774\ \cancel{\text{mol}_{N_2}} \times \frac{2\ \text{mol}_{NaN_3}}{3\ \cancel{\text{mol}_{N_2}}}$$
$$n_{NaN_3} = 1.849\ \text{mol}_{NaN_3}$$

**Step 5.** Determine the mass of sodium azide produced from the amount of sodium azide determined in Step 4 by multiplying this amount by a conversion factor derived from the molar mass of sodium azide.

$$M_{NaN_3} = 65.0 \text{ g/mol}$$
$$m_{NaN_3} = n_{NaN_3} \times \frac{65.0 \text{ g}}{1 \text{ mol}}$$
$$= 1.849\ \cancel{\text{mol}} \times \frac{65.0 \text{ g}}{1\ \cancel{\text{mol}}}$$
$$m_{NaN_3} = 120 \text{ g}$$

**Statement:** 120 g of sodium azide is required to produce 67.0 L of nitrogen gas at these conditions.

## Practice

SKILLS HANDBOOK A6

1. A 128 g sample of oxygen gas reacts completely with excess nitrogen monoxide according to the following balanced chemical equation:

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

   What volume of nitrogen dioxide will be produced at STP? T/I [ans: 179 L]
2. Sodium carbonate and hydrochloric acid react to give sodium chloride, carbon dioxide, and water. Calculate the mass of sodium carbonate that would be required to produce 2.00 L of carbon dioxide at 300.0 K and 101.3 kPa. T/I [ans: 8.61 g]
3. The following chemical equation shows the complete combustion of ethane:

   $2\ C_2H_6(g) + 7\ O_2(g) \rightarrow 6\ H_2O(g) + 4\ CO_2(g)$

   How many litres of oxygen would be required for the complete combustion of 82.0 L of ethane at 123 °C and 105 kPa? T/I [ans: 287 L]

**UNIT TASK BOOKMARK**

As you work on your Unit Task, on page 616, think about how small-scale methods of capturing carbon dioxide be scaled up to be used on an industrial level.

**Investigation 12.5.1**

**Gas Stoichiometry: Determining the Mass of Hydrogen Gas (p. 606)**
In this investigation you will set up a reaction that produces hydrogen gas. You will measure the volume of the gas, and the experimental conditions. You will use stoichiometry to determine the amount of gas produced. You will then combine these data to calculate the molar volume of hydrogen, and compare this value to the accepted molar volume.

## Mini Investigation

### Capturing Carbon Dioxide on a Small Scale

**Skills:** Planning, Performing, Observing, Analyzing, Communicating — SKILLS HANDBOOK A1, A2.4

Limewater, or calcium hydroxide solution, $Ca(OH)_2(aq)$, reacts to "capture" carbon dioxide gas. Carbon capture is a hot topic these days as global efforts unfold to control the emission of greenhouse gases. Many different technologies and solutions are currently being developed to sequester, capture, and convert carbon dioxide. Try your hand at carbon capture.

**Equipment and Materials:** chemical safety goggles; apron; plastic straw; small beaker containing limewater, $Ca(OH)_2(aq)$; dropper bottle containing bromothymol blue indicator

Do not suck up the limewater. If you get any limewater in your mouth, spit it into the sink immediately and rinse your mouth thoroughly.

1. Put on your chemical safety goggles and lab apron.
2. Place the plastic straw in the limewater. Gently exhale through the straw to bubble your breath into the limewater (**Figure 4**). Record your observations.
3. Dispose of your limewater according to your teacher's directions.
4. Obtain a fresh sample of limewater. Add 2 or 3 drops of bromothymol blue to the limewater and repeat Steps 2 and 3.

A. Write the balanced chemical equation for the reaction between calcium hydroxide and carbon dioxide. K/U C

B. What experimental evidence do you have to support your proposed chemical equation? T/I

C. Bromothymol blue is an acid–base indicator. Why did the colour change as carbon dioxide was added to the limewater? T/I

D. Would this be an effective method to remove carbon dioxide from the atmosphere? Explain. A

E. Design an experiment, based on this test, to determine the difference in the percentage of carbon dioxide in exhaled air compared with the percentage in atmospheric air.

**Figure 4** Always blow out through limewater.

## 12.5 Summary

- The law of combining volumes for gas reactions states that volumes of gaseous reactants and products of a chemical reaction are always in simple ratios of whole numbers, when measured at the same temperature and pressure. This can be used to solve simple gas stoichiometry problems.
- Stoichiometric principles and the use of the ideal gas law equation can be used to solve a variety of gas stoichiometry problems.

## 12.5 Questions

1. Butane gas, $C_4H_{10}$, undergoes complete combustion with excess oxygen to produce carbon dioxide and water vapour. T/I C
   (a) Write the equation for the total combustion of butane.
   (b) Determine the volume of carbon dioxide gas that would be produced by the complete combustion of 5.60 L of butane. Assume that all gases are at STP.
2. Potassium chlorate decomposes according to the following equation:

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

   What volume of a gas can be produced by the decomposition of 122.6 g of potassium chlorate measured under the following conditions? T/I
   (a) at STP
   (b) at SATP
3. Sodium metal with a total mass of 20.0 g is dropped into a beaker of water to produce hydrogen gas and sodium hydroxide. What volume of hydrogen gas will be produced if the temperature is 30 °C and the pressure is 125 kPa? T/I
4. Most cars today are powered by the combustion of gasoline. The main component of gasoline is octane, $C_8H_{18}(g)$, which burns as follows:

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

   One of the products is carbon dioxide, a major cause of climate change. T/I
   (a) Calculate the mass and volume of carbon dioxide gas produced by the combustion of a tank of gasoline. The mass of gasoline in the tank is 64.2 kg. The reaction takes place at a temperature of 80 °C and a pressure of 101.3 kPa. Assume that there is excess oxygen available and that only the complete combustion reaction occurs.
   (b) Plants effectively change carbon dioxide to glucose and oxygen gas through the process of photosynthesis. A single mature tree can absorb carbon dioxide at a rate of approximately 22 kg per year (**Figure 5**):

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

   Calculate the number of trees required to absorb, in one year, the carbon dioxide produced from one tank of gas (calculated in (a)).

**Figure 5** Trees provide many benefits, including carbon dioxide removal.

5. Calcium carbonate reacts with hydrochloric acid as follows:

   $CaCO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$ T/I
   (a) What volume of carbon dioxide, measured at 20 °C and 100.0 kPa, would be produced from 1025 g of calcium carbonate?
   (b) Calculate the mass of calcium carbonate required to produce 1550 L of carbon dioxide at 22.5 °C and 100.0 kPa.
6. Ammonium sulfate reacts with potassium hydroxide solution as follows:

   $(NH_4)_2SO_4(aq) + 2\ KOH(aq) \rightarrow 2\ NH_3(g) + K_2SO_4(aq) + 2\ H_2O(l)$

   Calculate the volume of ammonia gas, measured at 23 °C and 64 kPa, that could be produced from 264.0 g of ammonium sulfate and 280.0 g of potassium hydroxide. T/I
7. Propane gas burns in air:

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

   Determine the volume of carbon dioxide that will be formed when 54.0 g of propane, $C_3H_8$, reacts with excess oxygen at 25 °C and 202.1 kPa. T/I