Gas Law Stoichiometry

12.3

Many chemical reactions in everyday life involve gases. Figure 12.14 shows a common reaction that has a gas as a reactant. Other reactions, such as the electrolyzation of salt to give chlorine gas, have gases as products. To carry out an accurate and efficient reaction, scientists must know the number of moles of all the reactants. When one or more of the reactants is a gas, this means using the ideal gas law.

Figure 12.14 **In this photograph, oxygen gas reacts with calcium to produce calcium oxide.**

You have already learned that the ideal gas law can be used to solve for different variables in several different types of situations. As you may recall, *the term "stoichiometry" refers to the relationship between the number of moles of the reactants and the number of moles of the products in a chemical reaction*. In this section, you will learn how to use Gay-Lussac's law of combining volumes and the ideal gas law to solve stoichiometric problems that involve gases.

Section Preview/ Specific Expectations

In this section, you will

- **perform** stoichiometric calculations involving the number of moles, number of atoms, number of molecules, mass, and volume of substances in a balanced chemical reaction
- **determine** the molar volume of hydrogen in an investigation

Volume to Volume Stoichiometry

At the beginning of this chapter, you were introduced to Gay-Lussac's law of combining volumes: *When gases react, the volumes of the reactants and the products, measured at equal temperatures and pressures, are always in whole number ratios*. As well, you learned that the mole ratios from a chemical equation are the same as the ratios of the volumes of the gases.

This information will help you with a certain type of gas stoichiometry problem. When a gas reacts to produce another gas, you can use Gay-Lussac's law of combining volumes to find the volumes of the gases. The following Sample Problem shows you how.

CHECKPOINT

Go back to the beginning of this chapter. Make sure that you understand why the mole ratios are the same as the volume ratios.

Sample Problem

Gay-Lussac's Law of Combining Volumes

Problem

Ammonia is produced by a reaction of nitrogen gas and hydrogen gas. The chemical equation for the reaction is

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

Suppose that 12.0 L of nitrogen gas reacts with hydrogen gas at the same temperature and pressure.

(a) What volume of ammonia gas is produced?

(b) What volume of hydrogen is consumed?

What Is Required?

(a) Calculate the volume of ammonia gas produced when 12.0 L of nitrogen gas reacts.

(b) Calculate the volume of hydrogen gas used up by the reaction.

What Is Given?

From the equation, you know that 12.0 L of nitrogen gas is used. The mole ratios from the equation are

$$\frac{2 \text{ mol } NH_{3(g)}}{1 \text{ mol } N_{2(g)}}$$

$$\frac{3 \text{ mol } H_{2(g)}}{1 \text{ mol } N_{2(g)}}$$

Plan Your Strategy

You know that the mole ratios of the volumes of gases are the same as the ratios of the volumes. Therefore, you can use the mole ratios to find the volumes of ammonia gas and hydrogen gas. You do not need to use the temperature and pressure, since they remain the same in this problem.

Act on Your Strategy

(a) Let x be the volume of ammonia gas.

$$\frac{2 \text{ mol } NH_{3(g)}}{1 \text{ mol } N_{2(g)}} = \frac{x \text{ L } NH_{3(g)}}{12.0 \text{ L } N_{2(g)}}$$

$$(12.0 \text{ L } N_{2(g)})\frac{2 \cancel{\text{mol}} NH_{3(g)}}{1 \cancel{\text{mol}} N_{2(g)}} = \frac{x \text{ L } NH_{3(g)}}{12.0 \text{ L } N_{2(g)}}(12.0 \text{ L } N_{2(g)})$$

$$x = 24.0 \text{ L } NH_{3(g)}$$

Therefore, 24.0 L of ammonia gas is produced.

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(b) Let y be the volume of hydrogen gas.

$$\frac{3 \text{ mol } H_{2(g)}}{1 \text{ mol } N_{2(g)}} = \frac{y \text{ L } H_{2(g)}}{12.0 \text{ L } N_{2(g)}}$$

$$(12.0 \text{ L } N_{2(g)})\frac{3 \text{ mol } H_{2(g)}}{1 \text{ mol } N_{2(g)}} = \frac{y \text{ } H_{2(g)}}{12.0 \text{ L } N_{2(g)}}(12.0 \text{ L } N_{2(g)})$$

$y = 36.0 \text{ L } H_{2(g)}$

Therefore, 36.0 L of hydrogen gas is consumed.

Check Your Solution

The number of significant digits in the answer is the same as the number of significant digits in the question.

The mole ratio of ammonia to nitrogen is 2:1. Thus, it makes sense that the volume of ammonia gas is twice the volume of nitrogen gas.

The mole ratio of hydrogen to nitrogen is 3:1. It makes sense that the volume of hydrogen gas is three times the volume of nitrogen gas.

Practice Problems

25. Use the following balanced equation to answer the questions below.

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

(a) What is the mole ratio of oxygen gas to water vapour?

(b) What is the volume ratio of oxygen gas to water vapour?

(c) What is the volume ratio of hydrogen gas to oxygen gas?

(d) What is the volume ratio of water vapour to hydrogen gas?

26. 1.5 L of propane gas are burned in a barbecue. The following equation shows the reaction. Assume all gases are at STP.

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

(a) What volume of carbon dioxide gas is produced?

(b) What volume of oxygen is consumed?

27. Use the following equation to answer the questions below.

$$SO_{2(g)} + O_{2(g)} \rightarrow SO_{3(g)}$$

(a) Balance the equation.

(b) 12.0 L of sulfur trioxide, $SO_{3(g)}$, are produced at 100°C. What volume of oxygen is consumed?

(c) What assumption must you make to answer part (b)?

28. 2.0 L of gas A react with 1.0 L of gas B to produce 1.0 L of gas C. All gases are at STP.

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

(b) Each molecule of gas A is made of two identical "a" atoms. That is, gas A is really $a_{2(g)}$. In the same way, each molecule of gas B is made of two identical "b" atoms. What is the chemical formula of gas C in terms of "a" and "b" atoms?

Solving Gas Stoichiometry Problems

Earlier in this course, you learned how to do stoichiometry calculations. To solve gas stoichiometry problems, you will incorporate the ideal gas law into what you learned previously. The following steps will help you do this.

How to Solve Gas Stoichiometry Problems

1. Write a balanced equation for the reaction.
2. Write the given information under the appropriate reactants and products. Put a question mark under the reactant or product for which information is needed.
3. Convert all amounts to moles.
4. Compare molar amounts using stoichiometry ratios from the balanced equation. Solve for the unknown molar amount.
5. Convert the new molar amount into the units required. You may multiply by a conversion factor, or use a set of conditions with the ideal gas law, $PV = nRT$.

Using the Ideal Gas Law for the Gaseous Product of a Reaction

The best way to find out how to do a stoichiometry problem using the ideal gas law is to study an example. In the following Sample Problem, you will use a balanced equation and the ideal gas law to find the volume of a gas produced. (Refer to Chapter 4, section 4.1, if you want to review how to write balanced equations.)

Sample Problem

Mass to Volume Stoichiometry

Problem

Ancient alchemists liked to use strong sulfuric acid to produce dramatically dangerous effects. One interesting reaction occurs when sulfuric acid reacts with iron metal to produce gas and an iron(II) compound. What volume of gas is produced when excess sulfuric acid reacts with 40.0 g of iron at 18.0˚C and 100.3 kPa?

What Is Required?

Calculate the volume of gas that is produced when sulfuric acid reacts with iron under specific temperature and pressure conditions.

What Is Given?

Reactants: sulfuric acid and iron

Products: an iron(II) compound and a gas

Mass of iron = 40.0 g

Temperature = 18.0˚C

Pressure = 100.3 kPa

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Plan Your Strategy

Step 1 Write a balanced equation for the chemical reaction.

Step 2 Find the number of moles of iron present. Use this value, along with the mole ratios from the balanced equation, to find the number of moles of gas produced.

Step 3 Use the ideal gas law. You know the number of moles of gas, the temperature, and the pressure. (Do not forget to change the temperature to kelvins.) Solve for the volume of the gas.

Act on Your Strategy

Step 1 Write the balanced equation. (This reaction is a single displacement reaction.)

$$Fe_{(s)} + H_2SO_{4(aq)} \rightarrow H_{2(g)} + FeSO_{4(aq)}$$

$$40.0 \text{ g} \qquad\qquad ?\text{ L}$$

Step 2 Find the number of moles of iron, and the number of moles of gas.

To find the number of moles of iron, divide the mass by the molar mass. You can find the molar mass of iron in the periodic table: 55.85 g/mol. **Note:** If the reactant was a compound, such as $FeCl_2$, you would need to calculate the molar mass by adding the molar masses of all the atoms.

$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{40.0 \text{ g Fe}}{55.85 \text{ g/mol}} \\ &= 0.716 \text{ mol Fe} \end{aligned}$$

From the balanced equation, the mole ratio is

$$\frac{1 \text{ mol } H_2}{1 \text{ mol Fe}}$$

Use this ratio to find the number of moles of hydrogen gas formed by the reaction.

$$\frac{n \text{ mol } H_2}{0.716 \text{ mol Fe}} = \frac{1 \text{ mol } H_2}{1 \text{ mol Fe}}$$

$$(0.716 \text{ mol Fe}) \frac{n \text{ mol } H_2}{0.716 \text{ mol Fe}} = \frac{1 \text{ mol } H_2}{1 \text{ mol Fe}} (0.716 \text{ mol Fe})$$

$n = 0.716 \text{ mol } H_2$

Step 3 Use the ideal gas law to solve for the volume, since all the other quantities are now known.

First change the temperature to kelvins.

$18.0^\circ\text{C} + 273 = 291 \text{ K}$

You now have all the values you need to solve for volume.

$P = 100.3 \text{ kPa}$

$n = 0.716 \text{ mol}$

$R = 8.314 \text{ kPa·L/mol·K}$

$T = 291 \text{ K}$

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Web LINK

www.school.mcgrawhill/resources

One of the problems with air bags (see Figure 12.15 on the next page) is that they can harm a child or small adult because too much gas is produced. To find out what volume of nitrogen gas in an air bag is safe for a child, go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next. Use the volume you find to calculate the mass of $NaN_{3(s)}$ that is needed to produce this volume at 22.0°C and 105 kPa.

Figure 12.15 Air bags must be tested thoroughly before being manufactured for public use.

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$$PV = nRT$$
$$\therefore V = \frac{nRT}{P}$$
$$= \frac{0.716\ \cancel{\text{mol}} \times 8.314\ \cancel{\text{kPa}}\cdot\text{L}/\cancel{\text{mol}}\cdot\cancel{\text{K}} \times 291\ \cancel{\text{K}}}{100.3\ \cancel{\text{kPa}}}$$
$$= 17.3\ \text{L}$$

Therefore, 17.3 L of hydrogen gas is produced by this reaction at 18.0°C and 100.3 kPa.

Check Your Solution

The answer is slightly less than the molar volume of hydrogen gas at STP. Since less than one mole of hydrogen gas was formed, this seems reasonable.

Practice Problems

29. Engineers design automobile air bags that deploy almost instantly on impact. To do this, an air bag must provide a large amount of gas in a very short time. Many automobile manufacturers use solid sodium azide, NaN_3, along with suitable catalysts, to provide the gas that is needed to inflate the air bag. The balanced equation for this reaction is

$$2NaN_{3(s)} \rightarrow 2Na_{(s)} + 3N_{2(g)}$$

(a) What volume of nitrogen gas will be produced if 117.0 g of sodium azide are stored in the steering wheel at 20.2°C and 101.2 kPa?

(b) How many molecules of nitrogen are present in this volume?

(c) How many atoms are present in this volume?

30. 0.72 g of hydrogen gas, H_2, reacts with 8.0 L of chlorine gas, Cl_2, at STP. How many litres of hydrogen chloride gas, HCl, are produced?

31. How many grams of baking soda (sodium hydrogen carbonate, $NaHCO_3$), must be used to produce 45 mL of carbon dioxide gas at 190°C and 101.3 kPa in a pan of muffins? (The mole ratio of $NaHCO_3$ to CO_2 is 2:1.)

32. How much zinc (in grams) must react with hydrochloric acid to produce 18 mL of gas at SATP? (**Hint:** Zinc chloride, $ZnC\ell_{2(s)}$ is a product.)

33. 35 g of propane gas burned in a barbecue, according to the following equation:

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

All the gases are measured at SATP.

(a) What volume of water vapour is produced?

(b) What volume of oxygen is consumed?

34. What mass of oxygen is reacted to produce 0.62 L of water vapour at 100°C and 101.3 KPa? Start by balancing the following equation:

$$H_{2(g)} + O_{2(g)} \rightarrow H_2O_{(g)}$$

CHEM FACT

An automobile air bag fills up with about 65 L of nitrogen gas in approximately 27 ms. This can prevent a driver from being seriously injured. The sodium that is produced is extremely caustic, however. It reacts with iron(III) oxide as follows:

$$6Na_{(s)} + Fe_2O_{3(s)} \rightarrow 3Na_2O_{(s)} + 2Fe_{(s)}$$

The sodium oxide then reacts with carbon dioxide and water vapour.

$$Na_2O(s) + 2CO_{2(g)} + H_2O_{(g)} \rightarrow 2NaHCO_{3(s)}$$

The sodium hydrogen carbonate that is produced is a harmless substance. It is better known as baking soda.

Including Water Vapour Pressure in Gas Calculations

You can collect many gases by allowing them to bubble up through water into a container that is filled with water. (See Figure 12.16). This is the method your teacher used to collect the gas in Investigation 12-A. Unfortunately, molecules of water vapour mix with the gas sample. To avoid error, the pressure that was contributed by the water vapour must be subtracted when finding the pressure of the gas.

As an example, consider hydrogen gas, which is often collected over water. The hydrogen that is collected is a mixture of hydrogen and water vapour. As you learned from Dalton's law of partial pressures in Chapter 11, the pressure of this mixture is

$$P_{\text{total}} = P_{\text{hydrogen}} + P_{\text{water vapour}}$$

To find the partial pressure of dry hydrogen, subtract the pressure of the water vapour from the total pressure.

$$P_{\text{hydrogen}} = P_{\text{total}} - P_{\text{water vapour}}$$

The pressure of the water vapour is the same for any gas that is collected at a particular temperature. For example, the pressure of water vapour at 25°C is 3.17 kPa. Table 12.3 gives the pressure of water vapour at different temperatures.

When using the ideal gas law for a gas collected over water, you must correct the pressure before you substitute it into the gas law. The following Sample Problem shows you how to do this.

Figure 12.16 This is an efficient and convenient method for collecting hydrogen gas. Unfortunately molecules of water vapour mix with the gas sample.

Table 12.3
Pressure of Water Vapour

Temperature (°C)	Pressure (kPa)
17	1.94
18	2.06
19	2.20
20	2.34
21	2.49
22	2.64
23	2.81
24	2.98
25	3.17
26	3.36
27	3.56
28	3.78
29	4.00
30	4.24

Sample Problem

Calculating the Volume of a Gas Collected Over Water

Problem

A student reacts magnesium with excess dilute hydrochloric acid to produce hydrogen gas. She uses 0.15 g of magnesium metal. What volume of dry hydrogen does she collect over water at 28°C and 101.8 kPa?

What Is Required?

You need to find the volume of hydrogen collected over water in this reaction.

What Is Given?

$T = 28.0$°C

$P = 101.8$ kPa

Mass of magnesium (m) = 0.15 g

Pressure of water vapour at 28°C = 3.78 kPa

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Plan Your Strategy

Step 1 Write a balanced chemical equation for the reaction.

Step 2 Calculate the number of moles of magnesium by dividing the mass given (m) by the molar mass of magnesium (M). Use the number of moles of magnesium, along with the mole ratio from the equation, to calculate the number of moles of hydrogen gas produced by the reaction.

Step 3 Convert the temperature to kelvins. Since the hydrogen is collected over water, subtract the pressure of the water vapour at 28°C from the atmospheric pressure.

Step 4 Use the ideal gas law to find the unknown volume (V) of hydrogen gas.

Act on Your Strategy

Step 1 The balanced chemical equation is

$$\underset{0.15\ \text{g}}{Mg_{(s)}} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + \underset{?\ \text{L}}{H_{2(g)}}$$

Step 2 Find the number of moles of magnesium.

From the periodic table, the molar mass of magnesium is 24.31 g/mol.

$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{0.15\ \cancel{\text{g}}}{24.31\ \cancel{\text{g}}/\text{mol}} \\ &= 6.2 \times 10^{-3}\ \text{mol} \end{aligned}$$

The mole ratio of hydrogen gas to magnesium in this reaction is

$$\frac{1\ \text{mol } H_2}{1\ \text{mol Mg}}$$

Using the mole ratio,

$$\frac{n\ \text{mol } H_2}{6.2 \times 10^{-3}\ \text{mol}} = \frac{1\ \text{mol } H_2}{1\ \text{mol Mg}}$$

Cross multiply to get

$$(6.2 \times 10^{-3}\ \text{mol Mg})\ \frac{n\ \text{mol } H_2}{6.2 \times 10^{-3}\ \text{mol Mg}} = \frac{1\ \text{mol } H_2}{1\ \text{mol Mg}}\ (6.2 \times 10^{-3}\ \text{mol Mg})$$

$$n = 6.2 \times 10^{-3}\ \text{mol}$$

Step 3 Convert the temperature and pressure to kelvins.

$$\begin{aligned} T &= (28^\circ\text{C} + 273) \\ &= 301\ \text{K} \end{aligned}$$

The pressure that is exerted by the hydrogen gas is

$$\begin{aligned} P_{\text{hydrogen}} &= 101.8\ \text{kPa} - 3.78\ \text{kPa} \\ &= 98.0\ \text{kPa} \end{aligned}$$

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Step 4 Use the ideal gas law.

$$PV = nRT$$

$$\therefore V = \frac{nRT}{P}$$

$$= \frac{(6.2 \times 10^{-3}\ \cancel{\text{mol}})(8.314\ \cancel{\text{kPa}}\cdot\text{L}/\cancel{\text{mol}\cdot\text{K}})(301\ \cancel{\text{K}})}{(98.0\ \cancel{\text{kPa}})}$$

$$= 0.16\ \text{L}$$

The student collects 0.16 L of dry hydrogen.

Check Your Solution

The final answer is rounded to two significant digits. This is the least number of significant digits in the question. The mass of the magnesium is a small number. Therefore, the volume of the hydrogen produced is also a small number.

Using the Ideal Gas Law for the Gaseous Reactant of a Reaction

So far, you have used the ideal gas law for reactions with a gas *product*. You can also use the ideal gas law for reactions with a gas as a *reactant*. The following Sample Problem shows you how to do this.

Sample Problem

Space Shuttle Science: Gas as a Reactant

Problem

When astronauts travel in a space shuttle (Figure 12.17), carbon dioxide must be removed from the air they breathe. One method is to bubble the air in the shuttle though a solution of lithium hydroxide. The lithium hydroxide converts any carbon dioxide into lithium carbonate.

$$2LiOH_{(aq)} + CO_{2(g)} \rightarrow Li_2CO_{3(aq)} + H_2O_{(\ell)}$$

(You learned about a different method in Chapter 7 using solid LiOH.) Air containing 25.0 L of carbon dioxide is passed through 1.5 mol/L LiOH solution over a 20 min period. The atmospheric pressure in the shuttle is 0.85 atm, and the temperature is 28.3°C. What mass of lithium carbonate is produced?

Note: Although the air is bubbled through an aqueous solution, you do not need to consider the pressure of the water vapour. This is because you are dealing with the *reactant* not the *product*.

Figure 12.17

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What Is Required?

Find the mass of lithium carbonate that is produced when 25.0 L of carbon dioxide is bubbled through an aqueous solution of lithium hydroxide.

What Is Given?

$2LiOH_{(aq)} + CO_{2(g)} \rightarrow Li_2CO_{3(aq)} + H_2O_{(\ell)}$

1.0 L	$V = 25.0$ L
1.5 mol/L	$T = 28.3$°C
	$P = 0.85$ atm

Plan Your Strategy

Step 1 Convert the temperature to kelvins. Convert the pressure from atm to kPa, or use $R = 0.8206$ atm·L/mol·K.

Step 2 Use the ideal gas law to find the number of moles of carbon dioxide that reacts.

Step 3 Use the stoichiometry of the equation to determine the number of moles of lithium carbonate produced.

Step 4 Use the periodic table to determine the molar mass of lithium carbonate. To find the mass of lithium carbonate produced, multiply the number of moles by the molar mass.

Act on Your Strategy

Step 1 Convert the temperature and pressure.

$$\begin{aligned} T &= 28.3^\circ\text{C} + 273 \\ &= 301 \text{ K} \end{aligned}$$

$$0.85 \cancel{\text{atm}} \times \frac{101.3 \text{ kPa}}{1 \cancel{\text{atm}}} = 86 \text{ kPa}$$

Step 2 Find the number of moles of $CO_{2(g)}$.

$$PV = nRT$$

$$\begin{aligned} \therefore n &= \frac{PV}{RT} \\ &= \frac{(0.85 \cancel{\text{atm}})(25.0 \cancel{\text{L}})}{(0.08206 \cancel{\text{atm}}\cdot\cancel{\text{L}}/\text{mol}\cdot\cancel{\text{K}})(301 \cancel{\text{K}})} \\ &= 0.86 \text{ mol} \end{aligned}$$

Therefore, 0.86 mol of $CO_{2(g)}$ passes through the LiOH solution.

Step 3 Find the number of moles of Li_2CO_3 produced.

From the balanced equation, we know that 1 mol of carbon dioxide produces 1 mol of lithium carbonate.

$$\frac{n \text{ mol } Li_2CO_3}{0.86 \text{ mol } CO_2} = \frac{1 \text{ mol } Li_2CO_3}{1 \text{ mol } CO_2}$$

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$$(0.86 \text{ mol } CO_2)\,\frac{n \text{ mol } Li_2CO_3}{0.86 \text{ mol } CO_2} = \frac{1\text{mol } Li_2CO_3}{1\text{mol } CO_2}\,(0.86 \text{ mol } CO_2)$$

$n = 0.86$ mol of Li_2CO_3

Step 4 Find the molar mass of Li_2CO_3. Then find the mass of Li_2CO_3 produced.

$$\begin{aligned} M_{Li_2CO_3} &= (2 \times 6.94 + 12.01 + 3 \times 16.00) \text{ g/mol} \\ &= 73.89 \text{ g/mol} \end{aligned}$$

Thus, the mass of Li_2CO_3 produced is

$$\begin{aligned} m &= n \times M \\ &= 0.86 \text{ mol} \times 73.89 \text{ g/mol} \\ &= 64 \text{ g} \end{aligned}$$

Therefore, 64 g of Li_2CO_3 is produced in this reaction.

Check Your Solution.

The answer has two significant digits. This is the least number of significant digits in the question.

When the units cancel in the ideal gas equation, mol remains. When the units cancel in the final calculation, g remains.

The answer seems reasonable.

Practice Problems

35. Oxygen, O_2, reacts with magnesium, Mg, to produce 243 g of magnesium oxide, MgO, at 101.3 kPa and 45°C. How many litres of oxygen are consumed? Start by writing the balanced equation.

36. Zinc reacts with nitric acid to produce 34 L of dry hydrogen gas at 900 torr and 20°C. How many grams of zinc are consumed?

37. 0.75 L of hydrogen gas is collected over water at 25.0°C and 101.6 kPa. What volume will the dry hydrogen occupy at 103.3 kPa and 25.0°C?

38. 3070 kg of coal burns to produce carbon dioxide. Assume that the coal is 95% pure carbon and the combustion is 80% efficient. (**Hint:** The mole ratio of $C_{(s)}$ to $CO_{2(g)}$ is 5:4.) How many litres of carbon dioxide are produced at SATP?

39. When 7.48 g of iron reacts with chlorine gas, 21.73 g of product is formed.

(a) How many moles of chlorine are used?

(b) What is the formula for the product?

(c) Write the equation for the reaction that occurs.

Now you have a chance to do an exciting investigation: the reaction of magnesium metal with strong acid. Remember to take appropriate safety precautions and follow your teacher's directions when working with the acid.

Investigation 12-B

SKILL FOCUS

Predicting

Performing and recording

Analyzing and interpreting

The Production of Hydrogen Gas

In this investigation, you will produce hydrogen gas by reacting strong acid with magnesium metal.

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

You will collect the hydrogen gas over water in a graduated cylinder.

Question

What is the molar volume of dry hydrogen gas at STP? Calculate it using the volume of hydrogen gas that you produce and the mass of one reactant.

Prediction

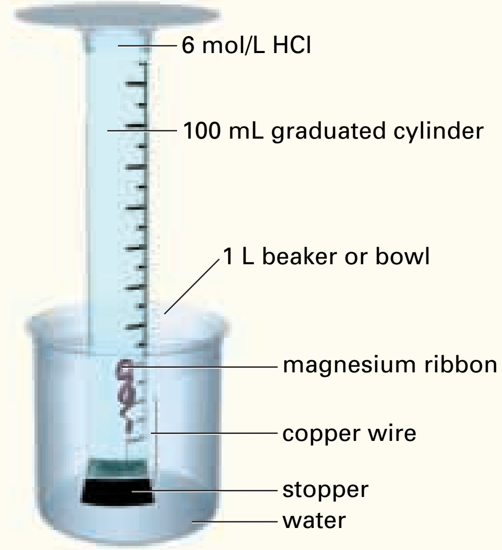
Predict the volume of hydrogen gas that will be produced. Use the mass of the magnesium and the balanced chemical equation given above. Assume 100% yield, and use regular stoichiometry. Organize your calculations clearly.

Safety Precautions

- Before beginning this investigation, check that there are no open flames (such as lit Bunsen burners) in the laboratory.
- The acid that you are using in this investigation is strong enough to burn. Wear your safety glasses and lab apron at all times. Handle the acid carefully. Wipe up any spills of water or acid immediately. If you accidentally spill any acid on your skin, wash it off immediately with large amounts of cool water.
- When you have finished the investigation, you can safely wash the products down the sink. You must dilute them, however, by running water down afterward.

Materials

scale or balance
100 mL graduated cylinder
stopper with two holes to fit graduated cylinder
1 L beaker or bowl
water at room temperature
6.0 mol/L hydrochloric acid, HCl
6 to 7 cm piece of magnesium ribbon
10 to 15 cm piece of copper wire
steel wool
barometer and thermometer
clamp and ring stand (optional)



Procedure

1. Prepare a table, in your notebook. Show your calculations.

Observations and Results

Observations	Trial 1	Trial 2 (if time permits)
mass of magnesium ribbon (g)		
temperature of water (°C)		
barometric pressure (kPa)		
volume of hydrogen collected (mL $\rightarrow$ L)		
vapour pressure of water at this temperature (kPa)*		

Results	Trial 1	Trial 2 (if time permits)
number of moles of magnesium (mol)		
volume of collected dry hydrogen at STP (L)		
molar volume of hydrogen at STP (L/mol)		

*Find the pressure of the water vapour in Table 12.3, or ask your teacher.

2. Obtain a piece of magnesium ribbon that is about 6 to 7 cm long. Use steel wool to clean the outside of the ribbon. Measure the mass of the ribbon.
3. Use the mass of the magnesium and the balanced equation to predict the volume of hydrogen gas that will be produced. Show your calculations in your notebook.
4. Fill the beaker (or bowl) about half full of water at room temperature. Measure the temperature of the water. (You will use this temperature to approximate the temperature of the gas produced.)
5. Measure and record the barometric pressure.
6. Add 15 mL of water to the graduated cylinder. Then, *very carefully*, pour 10 to 15 mL of 6 mol/L HCl into the graduated cylinder. *Very slowly and carefully*, pour water at room temperature down the *sides* of the cylinder until the cylinder is completely filled. Your objective in pouring the water this way is to avoid mixing it with the acid at the bottom of the cylinder. CAUTION Normally you should avoid adding water to an acid. Be particularly careful during this step.
7. Attach the magnesium ribbon to the copper wire. Dangle the magnesium in the graduated cylinder. The magnesium should hang 1 to 2 cm below the stopper. Put the stopper in the cylinder. Do not worry if a small amount of water overflows out of the cylinder.
8. Hold your gloved finger over the holes in the stopper. Tip the cylinder upside down into the 1 L beaker. Be careful that no air bubbles get into the cylinder. Hold or clamp the tube into place. Watch the reaction proceed. Record your observations in your notebook.
9. Add water, at room temperature, to the beaker until the level of the water inside the cylinder is exactly the same as the level of the water in the beaker. This equalizes the pressure of the hydrogen gas with the air pressure outside the tube. **Note:** Another way to equalize the pressure is to raise the graduated cylinder slightly to align the water levels.
10. Record the volume of the trapped gas.
11. All the magnesium should be used up by the reaction, since it is the limiting reagent. If any magnesium ribbon does remain after the reaction, rinse it with water, dry it with a paper towel, and measure its mass. To find the mass of the magnesium used up by the reaction, subtract the final mass from the initial mass.
12. Empty and clean all your apparatus. Clean your work space. Wash your hands.

Analysis

1. Calculate the molar volume of hydrogen. Use the volume of the H_2 gas, and the water temperature. Also, use the barometric pressure minus the pressure of the water vapour.
2. Use the combined or ideal gas law to translate the conditions to STP. Redo the calculations.

Conclusions

3. What was the class average for the molar volume of dry hydrogen gas at STP? How close was your molar volume to the class average?
4. How close was your molar volume to the accepted molar volume of a gas at STP? Calculate the percent error for your molar volume.
5. How would your results have been different had you not cleaned the magnesium ribbon before you used it?
6. What were some possible sources of error in your investigation?

Section Wrap-up

In this section, you learned how to use Gay-Lussac's law to calculate volumes of gases in a gas reaction. You also learned how to use the ideal gas law to find the volumes of gases used or produced in reactions. Building on Dalton's law of partial pressures, from Chapter 11, you learned how to calculate the molar volume of a gas collected over water. Finally, you learned, first-hand, how to produce hydrogen gas in a laboratory.

In the next section, you will see how gases in the atmosphere interact with the Sun's light. You will also find out about the dangers of gas pollution.

Web LINK

www.school.mcgrawhill.ca/resources/
What kind of company or industry uses, produces, or sells hydrogen gas? What is hydrogen gas used for? What safety precautions must be taken when handling it? To answer these questions, go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next.

Section Review

1. K/U Gay-Lussac's law of combining volumes provides a short-cut for some gas calculations. What type of gas reaction lets you use this law?

2. K/U What values should you measure in the laboratory in order to calculate the molar volume of a gas?

3. C Why is it necessary to correct the pressure of a gas that is collected over water? How would you do this?

4. K/U Which will occupy a greater volume: 2.0 mol of nitrogen gas at STP or 1.9 mol of oxygen gas at SATP? Explain.

5. I 2.0 L of hydrogen gas reacts with oxygen at 5˚C and 99 kPa. How many litres of water vapour are produced?

6. I 13 g of lead reacts with hydrofluoric acid to produce hydrogen gas at 22˚C and 88.3 kPa.
   - **(a)** Describe the apparatus you would need to carry out this reaction in the laboratory.
   - **(b)** How many litres of *dry* hydrogen gas are collected?

7. I Plants consume carbon dioxide gas as they produce sugar during photosynthesis.

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

   To produce 50 g of sugar, $C_6H_{12}O_6$, how many litres of gas at SATP does a sugar beet need to consume?

8. I 28 L of oxygen gas reacts with 52 L of hydrogen gas in an 80 L vessel at 25˚C and 3.0 atm.
   - **(a)** How many grams of water are produced? **Hint:** Find out which reagent will be used up by the reaction.
   - **(b)** What will the pressure be in the cylinder after the reaction if all of the water that is formed condenses and is removed?

9. MC After completing the Web Link task on this page, prepare a brochure for an imaginary hydrogen manufacturing company. Your brochure should advertise the many uses of your product. It must also contain safety information for the customer.