14.4 Gases: Mixtures and Movements

**Objectives**

14.4.1 Relate the total pressure of a mixture of gases to the partial pressures of the component gases.

14.4.2 Explain how the molar mass of a gas affects the rate at which the gas diffuses and effuses.

**Guide for Reading**

**Build Vocabulary**

Imagine a Picture As students learn about diffusion and diffusion, have them visualize each process at the microscopic level. Then have them describe what is happening.

**Reading Strategy**

Monitor Your Understanding If students are having difficulty reading Graham’s Law, have them think about how they are reading and how they might deal with the difficulty.

**INSTRUCT**

Connect to Your World

Have students study the photograph and read the text that opens the section. Ask students to try to explain why high-altitude climbers carry a supply of oxygen. (Students may recall that atmospheric pressure decreases with altitude. They may conclude that the drop in pressure is related to a decrease in the concentration of the gases in air, including oxygen.)

**Dalton’s Law**

Discuss

Assess students’ knowledge about gases by asking them to compare the number of particles in two identical, sealed containers of propane and helium gas at the same temperature and pressure. (Equal volumes of gases at the same temperature and pressure contain equal numbers of particles.)

**Table 14.1**

<table>
<thead>
<tr>
<th>Component</th>
<th>Volume (%)</th>
<th>Partial pressure (kPa)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nitrogen</td>
<td>78.08</td>
<td>79.11</td>
</tr>
<tr>
<td>Oxygen</td>
<td>20.95</td>
<td>21.22</td>
</tr>
<tr>
<td>Carbon dioxide</td>
<td>0.04</td>
<td>0.04</td>
</tr>
<tr>
<td>Argon and others</td>
<td>0.93</td>
<td>0.95</td>
</tr>
<tr>
<td><strong>Total</strong></td>
<td><strong>100.00</strong></td>
<td><strong>101.32</strong></td>
</tr>
</tbody>
</table>

This equation is a mathematical expression of a law proposed by Dalton. Dalton’s law of partial pressures states that, at constant volume and temperature, the total pressure exerted by a mixture of gases is equal to the sum of the partial pressures of the component gases.

\[ \text{Pressure}_{\text{total}} = \text{Pressure}_{1} + \text{Pressure}_{2} + \text{Pressure}_{3} + \ldots \]

**Connecting to Your World**

The top of Mount Everest is more than 29,000 feet above sea level. A list of gear for an expedition to Mount Everest includes climbing equipment such as an ice axe and a climbing harness. It includes ski goggles and a down parka with a hood. All the items on the list are important, but none is as important as the compressed-gas cylinders of oxygen. In this section, you will find out why a supply of oxygen is essential at higher altitudes.

**Dalton’s Law**

Gas pressure results from collisions of particles in a gas with an object. If the number of particles increases in a given volume, more collisions occur. If the average kinetic energy of the particles increases, more collisions occur. In both cases, the pressure increases. Gas pressure depends only on the number of particles in a given volume and on their average kinetic energy. Particles in a mixture of gases at the same temperature have the same average kinetic energy. So the kind of particle is not important.

Table 14.1 shows the composition of dry air. Partial pressure of nitrogen is 79.11 kPa. In a mixture of gases, the total pressure is the sum of the partial pressures of the gases.

**Section Resources**

- Guided Reading and Study Workbook, Section 14.4
- Core Teaching Resources, Section 14.4 Review
- Transparencies, T158–T159
- Laboratory Manual, Lab 25
- Small-Scale Chemistry Laboratory Manual, Lab 21
- Interactive Textbook with ChemASAP, Animation 17, 18; Problem-Solving 14.32; Assessment 14.4
- Go Online, Section 14.4
Model Partial Pressure

**Purpose** Students make a model that is analogous to Dalton’s law.

**Materials** balance, 4 marbles, 2 buttons, 3 pennies, 4 nickels

**Procedure** Have students find the total mass of the marbles and record it in a data table. Have them repeat this process for the buttons and the pennies. Then, have them add all the marbles, buttons, and pennies together and find the overall mass. Ask, How does the total mass of the items compare to the sum of the masses of the individual items? (The sum of the individual masses should equal the total mass.)

If the nickels are added to the mixture, how could you find the mass of the nickels? (The mass of the nickels could be calculated by finding a new total mass and subtracting from it the masses of the marbles, the buttons, and the pennies.)

**Expected Outcome** Students should recognize that mass is being used to represent pressure, and that the contribution of each type of item to the total mass is independent of the other types of items in the mixture.

**FYI**

You may want to compare the use of a compressed gas at high altitudes and the use of a compressed gas during deep underwater dives.

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**Relative Partial Pressure**

The relative partial pressure exerted by a gas in a mixture of gases does not vary with temperature, pressure, or volume of the mixture. The ratio of $n_1/n$ (the relative amount of the gas in the mixture) remains the same. Because the value of $n_1/n$ remains constant, the relative partial pressure ($P_1/P$) exerted by the gas remains the same.
Section 14.4 (continued)

Sample Problem 14.6

Answers
31. \(20.0 \text{kPa} + 46.7 \text{kPa} + 26.7 \text{kPa} = 93.4 \text{kPa} \) or \(9.34 \times 10^1 \text{kPa}\)
32. \(32.9 \text{kPa} - 6.6 \text{kPa} - 23.0 \text{kPa} = 3.3 \text{kPa}\)

Practice Problems Plus

The pressure in an automobile tire filled with air is 245.0 kPa. The \(P_{O_2} = 51.3 \text{kPa}, P_{CO_2} = 0.10 \text{kPa}, \) and \(P_{\text{others}} = 2.3 \text{kPa}\). What is the \(P_{N_2}\)? (191.3 kPa)

Use Visuals

Have students examine the visual on this page and read the caption. Ask, How does the amount of oxygen in the tank of compressed air compare to the amount of oxygen in the air you normally breathe? (Air is usually about 21% oxygen, so the air in the tanks contains about the same percentage of oxygen.)

Use Visuals

Solve Problem 32 with the help of an interactive guided tutorial.

Sample Problem 14.6

Using Dalton's Law of Partial Pressures
Air contains oxygen, nitrogen, carbon dioxide, and trace amounts of other gases. What is the partial pressure of oxygen \(P_{O_2}\) at 101.30 kPa of total pressure if the partial pressures of nitrogen, carbon dioxide, and other gases are 79.10 kPa, 0.040 kPa, and 0.94 kPa, respectively?

1. Analyze
Get the answers and the unknown.

<table>
<thead>
<tr>
<th>Knowns</th>
<th>Unknown</th>
</tr>
</thead>
<tbody>
<tr>
<td>(P_{N_2} = 79.10 \text{kPa})</td>
<td>(P_{O_2})</td>
</tr>
<tr>
<td>(P_{CO_2} = 0.040 \text{kPa})</td>
<td></td>
</tr>
<tr>
<td>(P_{\text{others}} = 0.94 \text{kPa})</td>
<td></td>
</tr>
</tbody>
</table>

2. Calculate
Solve for the unknown.
Rearrange Dalton’s law to isolate \(P_{O_2}\). Substitute the values for the partial pressures and solve the equation.

\[
P_{O_2} = P_{\text{total}} - (P_{N_2} + P_{CO_2} + P_{\text{others}})
\]

\[
P_{O_2} = 101.30 \text{kPa} - (79.10 \text{kPa} + 0.040 \text{kPa} + 0.94 \text{kPa})
\]

\[
P_{O_2} = 21.22 \text{kPa}
\]

3. Evaluate
Does the result make sense?
The partial pressure of oxygen must be smaller than that of nitrogen because \(P_{\text{total}}\) is only 101.30 kPa. The other partial pressures are small, so an answer of 21.22 kPa seems reasonable.

Practice Problems

31. Determine the total pressure of a gas mixture that contains oxygen, nitrogen, and helium. The partial pressures are: \(P_{O_2} = 20.0 \text{kPa}, P_{N_2} = 46.7 \text{kPa}, \) and \(P_{He} = 26.7 \text{kPa}\).

32. A gas mixture containing oxygen, nitrogen, and carbon dioxide has a total pressure of 32.9 kPa. If \(P_{O_2} = 6.6 \text{kPa}\) and \(P_{N_2} = 23.5 \text{kPa}\), what is \(P_{CO_2}\)?
**Graham's Law**

Suppose you open a perfume bottle in one corner of a room. At some point a person standing in the opposite corner will be able to smell the perfume. Molecules in the perfume evaporate and diffuse, or spread out, through the air in the room. Diffusion is the tendency of molecules to move toward areas of lower concentration until the concentration is uniform throughout. In Figure 14.18A, bromine vapor is diffusing through the air in a graduated cylinder. The bromine vapor in the bottom of the cylinder has started to move upward toward the area where there is a lower concentration of bromine. In Figure 14.18B, the bromine has diffused almost to the top of the cylinder. If the process is allowed to continue, the bromine vapor will spill out of the cylinder.

There is another process that involves the movement of molecules in a gas. This process is called effusion. During effusion, a gas escapes through a tiny hole in its container. With effusion and diffusion, the type of particle is important. Gases of lower molar mass diffuse and effuse faster than gases of higher molar mass.

**Thomas Graham's Contribution**

The Scottish chemist Thomas Graham studied rates of effusion during the 1840s. From his observations, he proposed a law. Graham's law of effusion states that the ratio of the rate of effusion of a gas is inversely proportional to the square root of the gas’s molar mass. This law can also be applied to the diffusion of gases.

Graham’s law makes sense if you know how the mass, velocity, and kinetic energy of a moving object are related. The expression that relates the mass \( m \) and the velocity \( v \) of an object to its kinetic energy \( KE \) is

\[
KE = \frac{1}{2}mv^2.
\]

For the kinetic energy to be constant, any increase in mass must be balanced by a decrease in velocity. For example, a ball with a mass of 2 g must travel at 5 m/s to have the same kinetic energy as a ball with a mass of 1 g traveling at 7 m/s. There is an important principle here: If two objects with different masses have the same kinetic energy, the lighter object must move faster.

**Word Origins**

Diffusion and effusion come from the Latin fundere meaning “to pour.” They differ only in their prefixes. The prefix dis means “apart.” The prefix ex means “out.” How do these prefixes help to contrast what happens to a gas during diffusion and effusion?

**Effusion**

**Purpose**

Students compare effusion rates of helium and air.

**Materials**

2 round, identical balloons; helium; metric tape measure.

**Procedure**

Fill two identical, round non-latex balloons with equal volumes of helium and air. Have volunteers measure the circumference of the balloons. One day later, measure the circumferences again.

**Expected Outcome**

Helium effused from the balloon at a faster rate.

**Class Activity**

**Download a worksheet on Diffusion/Effusion for students to complete, and find additional teacher support from NSTA SciLinks.**

**Answers to...**

Figure 14.18 The bromine will spill out of the container.

The Behavior of Gases 435


Section 14.4 (continued)

**Assess**

Evaluate Understanding  
L2

Tell students that the partial pressures of oxygen and hydrogen gases in a container are both 100 kPa. Ask, in which sample are there more molecules? In which sample do the molecules have greater average kinetic energy? Both gases have the same number of molecules and same average kinetic energy. Ask, About how much faster does helium diffuse compared to oxygen? (Helium diffuses almost three times faster than oxygen.)

Reteach

Remind students that diffusion is a general term that applies to molecules moving away from a region of high concentration. Effusion is a specific example of diffusion in which molecules pass through a narrow opening. Graham’s law applies to both.

Both tables report data for dry air. Both show that air is mainly nitrogen and oxygen. Table 14.1 uses percents and includes a compound. The table on R4 uses ppm and provides specific data for more elements.

If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 14.4.

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Section 14.4 Assessment

33. Total pressure is equal to the sum of the partial pressures of the components.

34. Gases with lower molar masses diffuse and effuse faster than gases with higher molar masses.

35. Subtract the partial pressures of the other gases from the total pressure.

36. During effusion, a gas escapes through a tiny hole in its container. In both cases, the rate depends on the molar mass.

37. The rate of effusion of two gases in a mixture is inversely proportional to the square roots of their molar masses.

38. Carbon monoxide and nitrogen have almost identical molar masses when the masses are rounded to two significant figures (28 g).

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Comparing Effusion Rates  
L1

The balloons in Figure 14.19 is used in holiday parades. It is filled with helium so that it will float above the crowd along the parade route. There is a drawback to using helium in a balloon. Both helium atoms and the molecules in air can pass freely through the tiny pores in a balloon. But a helium-filled balloon will deflate sooner than an air-filled balloon. Kinetic theory can explain this difference.

If the balloons are at the same temperature, the particles in each balloon have the same average kinetic energy. But helium atoms are less massive than oxygen or nitrogen molecules. So the molecules in air move more slowly than helium atoms with the same kinetic energy. Because the rate of effusion is related only to a particle’s speed, Graham’s law can be written as follows for two gases, A and B.

The rates of effusion of two gases are inversely proportional to the square roots of their molar masses. You can use the expression to compare the rates of effusion of nitrogen (molar mass \(28.0\) g) and helium (molar mass \(4.0\) g). Helium effuses (and diffuses) nearly three times faster than nitrogen at the same temperature.

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\[
\frac{\text{Rate}_A}{\text{Rate}_B} = \sqrt{\frac{M_B}{M_A}} = \frac{7.0}{2.7} = 2.7
\]

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Section 14.4 Assessment

436 Chapter 14
Small-Scale Lab

**Diffusion**

**Purpose**
To infer diffusion of a gas by observing color changes during chemical reactions.

**Materials**
- clear plastic cup or Petri dish
- reaction surface
- dropper bottles containing bromthymol blue, hydrochloric acid, and sodium hydrogen sulfite
- ruler
- cotton swab
- NaOH, NH₄Cl, KI, and NaHSO₃ (optional)

**Procedure**
1. Use the plastic cup or Petri dish to draw the large circle shown below on a sheet of paper.

   ![Small drops are BTB](image)

2. Place a reaction surface over the grid and add small drops of bromthymol blue (BTB) in the pattern shown by the small circles. Make sure the drops do not touch one another.

3. Mix one drop each of hydrochloric acid (HCl) and sodium hydrogen sulfite (NaHSO₃) in the center of the pattern.

4. Place the cup or Petri dish over the grid and observe what happens.

5. If you plan to do You’re The Chemist Activity 1, don’t dispose of your materials yet.

**Analyze and Conclude**
1. The drops near the center change immediately. As the gas diffuses, all the drops change color. The color change begins at the outer edge of each drop.

   center yellow

2. Draw a series of pictures showing how one of the BTB drops might look over time if you could view the drop from the side.

3. The BTB changed even though you added nothing to it. If the mixture in the center circle produced a gas, would this explain the change in the drops of BTB? Use kinetic theory to explain your answer.

4. Translate the following word equation into a balanced chemical equation:
   \( \text{NaNO}_2 + \text{HCl} \rightarrow \text{SO}_2 + \text{H}_2\text{O} + \text{NaCl} \)

**You’re The Chemist**
This follows small-scale activities allow you to develop your own procedures and analyze the results.

1. **Analyze**
   - Carefully absorb the center mixture of the original experiment onto a cotton swab and replace it with hydrochloric acid at the center. Record your results.

2. **Design**
   - Design an experiment to observe the effect of the size of the BTB drops on the rate at which they change.
   - Explain your results in terms of kinetic theory.

3. **Repeat**
   - Repeat the original experiment, using KI in place of BTB and mixing sodium nitrite (NaNO₂) with hydrochloric acid at the center. Record your results.
   - Write and balance an equation for the reaction: NaHSO₃ reacts with HCl to produce sulfur dioxide gas, water, and sodium chloride.

**Key**
Center mixing is HCl + NaHSO₃

**Safeguards**
Always slowly add acid to water.

**Teaching Tips**
- Before class, add a drop of BTB to the paper to be sure the paper is not acidic. The BTB should remain blue.

**Expected Outcome**
The drops of BTB turn yellow, starting with those closest to the center.

**You’re the Chemist!**
1. As ammonia diffuses, BTB changes from yellow to blue.
   - \( \text{NH}_4\text{Cl} + \text{NaOH} \rightarrow \text{NH}_3 + \text{H}_2\text{O} + \text{NaCl} \)

2. Vary the size of the BTB drops from “pin-heads” to “puddles.” Tiny drops are better able to detect small quantities of gases.

3. The KI turned orange in the same manner as the BTB turned yellow.
   - \(3\text{NaNO}_2 + 2\text{HCl} \rightarrow 2\text{NO} + \text{H}_2\text{O} + \text{NaNO}_3 + 2\text{NaCl} \)

**For Enrichment**
Have students design an experiment that shows the effect of temperature on rate of diffusion. The drops change color more quickly at higher temperatures.