**14.4 Gases: Mixtures and Movements**

**FOCUS**

**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 *effusion* 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**

**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, air that does not contain any water vapor. The contribution each gas in a mixture makes to the total pressure is called the *partial pressure* exerted by that gas. In dry air, the 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.

\[
P_{\text{total}} = P_1 + P_2 + P_3 + \ldots
\]

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.

**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>Total</td>
<td>100.00</td>
<td>101.32</td>
</tr>
</tbody>
</table>

**Section Resources**

**Print**

- *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*

**Technology**

- *Interactive Textbook with ChemASAP, Animation 17, 18; Problem-Solving 14.32; Assessment 14.4*
- *Go Online, Section 14.4*
Look at Figure 14.16. Containers A, B, C, and T (for total) have the same volume and are at the same temperature. The gases in containers A, B, and C are combined in container T. Because the volumes are identical, each gas in the mixture exerts the pressure it exerted before the gases were mixed in container T. So the pressure in container T (550 kPa) is the sum of the pressures in containers A, B, and C (100 + 250 + 200 kPa).

If the percent composition of a mixture of gases does not change, the fraction of the pressure exerted by a gas does not change as the total pressure changes. This fact is important for people who must operate at high altitudes. For example, at the top of Mount Everest, the total atmospheric pressure is 33.73 kPa. This is about one-third of its value at sea level. The partial pressure of oxygen is also reduced by one third, to 7.06 kPa. The partial pressure of oxygen must be 10.67 kPa or higher to support respiration in humans. The climber in Figure 14.17 needs an oxygen mask and a cylinder of compressed oxygen to survive.

**Facts and Figures**

**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.

**Answers to...**

Figure 14.16 The ratio of the particles in Containers C and A (2:1) is equal to the ratio of their partial pressures.
Sample Problem 14.6

Answers
31. 20.0 kPa + 46.7 kPa + 26.7 kPa = 93.4 kPa or 9.34 × 10^1 kPa
32. 32.9 kPa − 6.6 kPa − 23.0 kPa = 3.3 kPa

Practice Problems Plus
The pressure in an automobile tire filled with air is 245.0 kPa. The $P_{O_2} = 51.3$ kPa, $P_{CO_2} = 0.10$ kPa, and $P_{others} = 2.3$ 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.)

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 List the knowns and the unknown.
   - Knowns
     - $P_{N_2} = 79.10$ kPa
     - $P_{CO_2} = 0.040$ kPa
     - $P_{others} = 0.94$ kPa
     - $P_{total} = 101.30$ kPa
   - Unknown
     - $P_{O_2}$

   Use Dalton’s law of partial pressures ($P_{total} = P_{O_2} + P_{N_2} + P_{CO_2} + P_{others}$) to calculate the unknown value ($P_{O_2}$).

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.
   
   \[
   P_{O_2} = P_{total} - (P_{N_2} + P_{CO_2} + P_{others})
   \]
   
   \[
   = 101.30 \text{ kPa} - (79.10 \text{ kPa} + 0.040 \text{ kPa} + 0.94 \text{ kPa})
   \]
   
   \[
   = 21.22 \text{ kPa}
   \]

3. Evaluate Does the result make sense?
   The partial pressure of oxygen must be smaller than that of nitrogen because $P_{total}$ 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: $P_{O_2} = 20.0$ kPa, $P_{N_2} = 46.7$ kPa, and $P_{He} = 26.7$ kPa.
32. A gas mixture containing oxygen, nitrogen, and carbon dioxide has a total pressure of 32.9 kPa. If $P_{O_2} = 6.6$ kPa and $P_{N_2} = 23.0$ kPa, what is $P_{CO_2}$?
The Behavior of Gases

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 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 = 1/2mv^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.

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?

Word Origins

During diffusion, the particles move apart. During effusion, the particles move out of a tiny hole. (Students may have trouble with the scientific definitions of effusion and diffusion because of the multiple, often overlapping, meanings of these terms in common usage. For example, in many dictionaries one definition of effuse is “to spread out; diffuse; radiate.”)

Use Visuals

Figure 14.18 Have students study the photograph. Explain that diffusion is a relatively slow process whose rate is inversely proportional to the square root of the molar mass of the gas.

CLASS Activity

Effusion

Purpose Students compare effusion rates of helium and air.

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

Safety Wear safety goggles when handling balloons.

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.

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.

Differentiated Instruction

Less Proficient Readers Have students work in groups to rank the relative rates of diffusion at a constant temperature for carbon dioxide, helium, and nitrogen. (helium > nitrogen > carbon dioxide)
Section 14.4 (continued)

ASSESS

Evaluate Understanding

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.

AsseSS

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Elements Handbook

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.

Interactive Textbook

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

Section 14.4 Assessment

33. **Key Concept** In a mixture of gases, how is the total pressure determined?

34. **Key Concept** What is the effect of molar mass on rates of diffusion and effusion?

35. How is the partial pressure of a gas in a mixture calculated?

36. What distinguishes effusion from diffusion? How are these processes similar?

37. How can you compare the rates of effusion of two gases in a mixture?

38. Explain why the rates of diffusion of nitrogen gas and carbon monoxide are almost identical at the same temperature.

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Comparing Effusion Rates

The balloon 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.

$$\frac{\text{Rate}_A}{\text{Rate}_B} = \sqrt{\frac{\text{molar mass}_B}{\text{molar mass}_A}}$$

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.

$$\frac{\text{Rate}_{He}}{\text{Rate}_{N_2}} = \sqrt{\frac{28.0 \text{ g}}{4.0 \text{ g}}} = \sqrt{7.0} = 2.7$$

Table 14.1 on page 432 and the Elements in the Atmosphere table on page R4 provide data on the composition of air. Look at the data included in each table. Identify two ways in which the tables are similar. Describe at least three differences.
**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 NaNO₂ (optional)

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

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 the changes. Explain your results in terms of kinetic theory.

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

**Analyze and Conclude**
Using your experimental data, record the answers to the following questions below your data table.

1. Describe in detail the changes you observed in the drops of BTB over time. Draw pictures to illustrate the changes.

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: Sodium hydrogen sulfite reacts with hydrochloric acid to produce sulfur dioxide gas, water, and sodium chloride.

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

1. **Analyze It!** Carefully absorb the center mixture of the original experiment onto a cotton swab and replace it with hydrochloric acid to produce sulfur dioxide gas, water, and sodium chloride.

2. **Design It!** Design an experiment to observe the effect of temperature on the rate of diffusion. The drops change color more quickly at higher temperatures.

3. The KI turned orange in the same manner as the BTB turned yellow. Write and balance a chemical equation to describe this reaction.

4. **Analyze It!** 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. NaNO₂ reacts with HCl to produce nitrogen monoxide gas, water, sodium nitrate, and sodium chloride.

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

**Safety**
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.

**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.