SECTION 11.2

Gases and Pressure Changes

If you have flown in an airplane, you may have felt discomfort in your ears during take-off or landing. As soon as your ears "popped," they probably felt better. A similar experience is common while riding up and down on an elevator. Your ears become blocked and then unblocked due to changes in atmospheric pressure. Although you are usually unaware of the effect of atmospheric pressure on your body, the atmosphere is always exerting a large amount of pressure on you from all directions.

To describe and explain the behaviour of the gases in Earth's atmosphere, as well as the behaviour of other gases, you need to understand the meanings of pressure and gas pressure. *Pressure* is the force that is exerted on an object per unit of surface area. The phrase "per unit of surface area" means the area over which the force is distributed. The equation for pressure is

$$\text{pressure} = \frac{\text{force}}{\text{area}} \quad \text{or} \quad P = \frac{F}{A}$$

The SI unit for force is the newton (N), and the unit for area is the square metre (m^2). Therefore, the corresponding unit of pressure is newtons per square metre (N/m^2). Later in this chapter, you will learn how this unit for pressure is related to other commonly used units for presssure, such as the pascal (Pa) and millimetres of mercury (mmHg).

Key Terms

atmospheric pressure
standard atmospheric pressure (SAP)
Boyle's law

Atmospheric Pressure

Earth's atmosphere is a spherical envelope of gases that surrounds the planet and extends from Earth's surface outward to space. The gas molecules that make up the atmosphere are pulled down toward Earth's surface by gravity, and these molecules exert pressure on all objects on Earth. Thus, the atmosphere exerts pressure on everything on Earth's surface. Earth's **atmospheric pressure** may be described as the force that a column of air exerts on Earth's surface, divided by the area of Earth's surface at the base of the column, as shown in **Figure 11.6**. The force that the column of air exerts is its weight. Unlike force, however, which is exerted in only one direction, pressure is exerted *equally in all directions*.

atmospheric pressure the force exerted on Earth's surface by a column of air over a given area

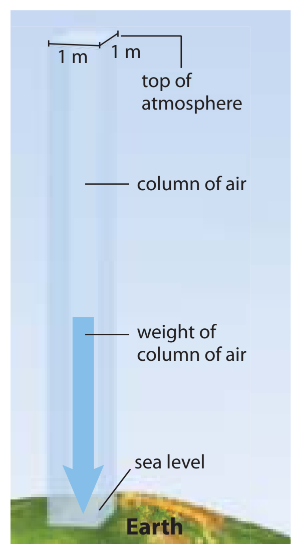

Figure 11.6 A column of air extending from sea level to the top of the atmosphere, with a cross-sectional area of 1 m^2, weighs 101 325 N. The mass of this column of air is 10 329 kg.

Early Studies of Atmospheric Pressure

People invented technologies that made use of atmospheric pressure before anyone understood how these technologies worked! For example, in 1594, the Italian scientist Galileo Galilei (1564–1642) was awarded a patent for his invention of a suction pump that used air to lift water up to the surface from about 10 m underground. (No plans or illustrations of this pump exist, but the pump operated using the kinetic energy supplied by one horse.) Although Galileo's pump functioned well, nobody—including Galileo himself—understood how water moved up the tube of the pump and why the pump could lift the water no higher than 10 m.

From 1641 to 1642, Evangelista Torricelli (1608–1647) served as Galileo's assistant. His work eventually led to an understanding of how the water moved up the tube of the pump. Torricelli hypothesized that the water rose in the tube because the surrounding air was pushing down on the rest of the water. Instead of water, he used mercury for his studies, because mercury is 13.6 times more dense than water. Thus, Torricelli could use a column of mercury that is 13.6 times shorter than the 10 m column of water.

Torricelli's Hypothesis Is Confirmed

Torricelli designed an experiment with an apparatus like the one in **Figure 11.7**. He proposed that if a long tube was filled with mercury and inverted into a dish of mercury, the mercury in the tube would drop down and leave a vacuum at the closed end of the tube. Torricelli hypothesized that the pressure that the column of mercury would exert on the mercury in the dish would be equal to the pressure that the atmosphere was exerting on the surface of the mercury in the dish outside the tube. His hypothesis was verified by experiments, and the limitations of the early suction pump could now be explained. The early suction pump could not lift water more than 10 m because atmospheric pressure was *pushing* the water up, not because the pump was pulling the water up. Atmospheric pressure is about approximately the same as the pressure exerted by 10 m of water—thus the 10 m limit.

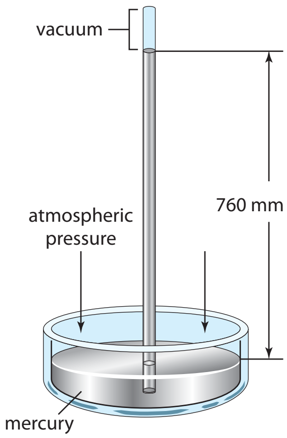

Figure 11.7 Torricelli's apparatus used mercury to test the hypothesis that underground water was being pushed up the tube of a water pump because of the air pressure acting on the surrounding water.

Predict *What would cause the level of mercury in the tube to change?*

The Relationship between Atmospheric Pressure and Altitude

In 1647, the French scientist and philosopher, Blaise Pascal (1623–1662), read a letter written by Torricelli, in which he compared the atmosphere to an ocean of air. In the letter, Torricelli hypothesized that the weight of air might be greater near Earth's surface than it was at the top of mountains. The next year, Pascal designed an experiment to test Torricelli's hypothesis that atmospheric pressure decreases with altitude (distance from Earth's surface). He asked his brother-in-law, Florin Perier, to carry an apparatus like Torricelli's up and down a mountain called Puy de Dôme, shown in **Figure 11.8**. Perier measured the length of the column of mercury at the base of the mountain, during his climb of the mountain, and on the top. Perier verified that as he ascended the mountain, the column of mercury became shorter. At the top of the mountain, the column of mercury was 76 mm shorter than it was at the base of the mountain.

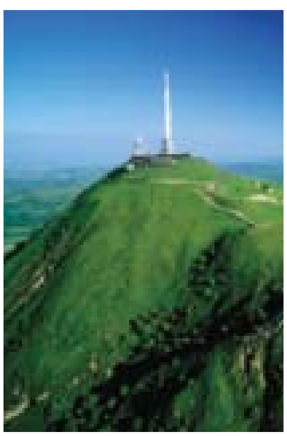

Figure 11.8 Puy-de-Dôme is a volcanic mountain that is the highest mountain in south-central France. Today, modern communications equipment is located on the top of this mountain.

Figure 11.9 shows that at higher altitudes (distances from Earth's surface) the atmospheric pressure is lower and the density of air particles is less than at lower altitudes. Why is this the case? Each layer of the atmosphere exerts a force on the layer below it. Because the weight of the entire column of air exerts a force on the bottom layer, the bottom layer is the most compressed. As the altitude increases, the amount of air above that level becomes smaller and, therefore, exerts a smaller force on the air just below it. The higher layers of air are less compressed than the lower layers.

Besides demonstrating that atmospheric pressure exists and that it changes with altitude, the combined efforts of Torricelli, Pascal, and Perier resulted in an instrument for measuring atmospheric pressure: the mercury barometer. Barometers based on Torricelli's design have been in use since the mid-1600s. Although newer technologies have been developed, many mercury barometers are still used around the world today.

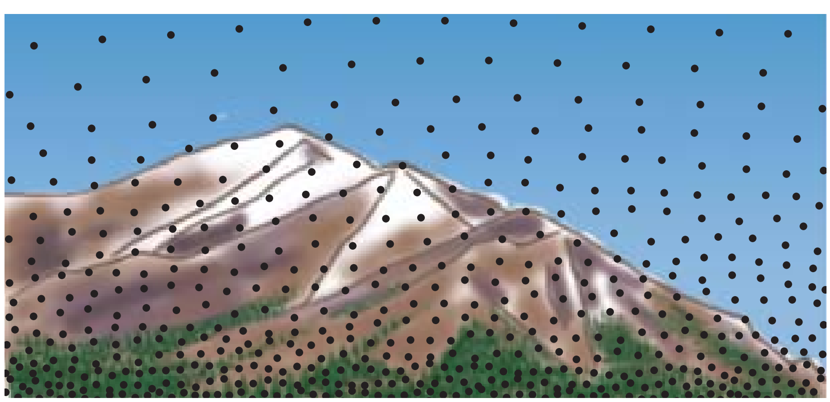

Figure 11.9 In this diagram, the dots represent air molecules. People often refer to the air "thinning" at increased altitudes. This means that there are fewer gas molecules in the air for a given volume at lower atmospheric pressure.

Units of Gas Pressure

Because mercury barometers were often used to measure atmospheric pressure, a common unit of pressure is the millimetre of mercury, or mmHg. **Standard atmospheric pressure (SAP)**, the atmospheric pressure in dry air at a temperature of 0°C at sea level, is 760 mmHg. Since standard atmospheric pressure is a common reference point, the unit *atmosphere* (atm) is also often used.

standard atmospheric pressure (SAP) atmospheric pressure in dry air at a temperature of 0°C at sea level

Recall that the newton per square metre (N/m^2) is the SI unit for pressure. This unit is also called the pascal (Pa), in honour of Blaise Pascal. Standard atmospheric pressure is 101 325 Pa. Because this is such a large number, the unit kPa (kilopascal) is commonly used. Thus, standard atmospheric pressure is often expressed as 101.325 kPa.

Another unit often used in chemistry is the bar (b). One bar is equal to 100 kPa, and 1 atm is equal to 1.01325 bar. In the United States, pressure is often measured and expressed in pounds per square inch (psi). You may have seen the recommended pressure for bicycle and automotive tires in units of psi. Finally, in honour of Torricelli's work, standard atmospheric pressure has also been defined as 760 torr. One torr represents a column of mercury that is 1 mm high at 0°C. The following expression summarizes the various units of pressure and their equivalent values.

$$1 \text{ atm} = 760 \text{ mmHg} = 760 \text{ torr} = 101\ 325 \text{ Pa} = 101.325 \text{ kPa} = 1.01325 \text{ bar} = 14.7 \text{ psi}$$

Different products and technologies have tended to use or report atmospheric pressure in different units. **Table 11.3** lists some of the instruments that use these different units to measure pressure.

Table 11.3 Units of Pressure Used for Various Instruments

Unit of Pressure	Symbol	Examples of Instruments That Use the Unit
standard atmosphere	atm	Gas compressors, pneumatic tools (tools such as jackhammers driven by compressed gas)
millimetres of mercury	mmHg	Blood pressure meters, barometers
torr	torr	Vacuum pumps
pascal	Pa	Pressure sensors in pipelines
kilopascal	kPa	Tire inflation gauges; heating, ventilating, and air-conditioning systems
bar	bar	Pressure sensors in scuba gear, steam traps used to remove condensed water from pipes carrying steam
millibar	mb	Barometers
pounds per square inch	psi	Hydraulic pumps, tire inflation gauges

Converting among Units of Pressure

Because people in different industries report pressure using different units, it is often necessary to convert between different units of pressure. Knowing the equivalent unit values makes the conversion straightforward. For example, suppose that the atmospheric pressure in Kenora, Ontario, is measured to be 732 mmHg and you want to know what this pressure is in kilopascals. Because 760 mmHg is equivalent to 101.325 kPa, conversion of 732 mmHg to kPa is

$$732 \cancel{\text{mmHg}} \times \left(\frac{101.325 \text{ kPa}}{760 \cancel{\text{mmHg}}}\right) = 97.6 \text{ kPa.}$$

Learning Check

7. What is atmospheric pressure?
8. Explain how Torricelli's apparatus worked.
9. Convert each of the following to the indicated unit.
   a. 3.58 atm to kPa
   b. 20.5 psi to atm
   c. 770 mmHg to kPa
   d. 470 torr to Pa
10. If the optimum tire pressure for a bicycle is 3 bar, and your tire pressure gauge is in units of psi, develop a formula that you can use to convert these units.
11. Why must mountain climbers understand the relationship between altitude and atmospheric pressure?
12. To make a birdbath, you fill a 2 L soft-drink bottle with water and invert it in a dish of water. When the level of the water in the dish falls below the level of the water at the rim of the bottle, water flows from the bottle to refill the dish. Explain why this happens.

The Relationship between Gas Pressure and Volume

Meteorologists use weather balloons to carry instrument packages called radiosondes high into the atmosphere. The balloon is partly inflated with helium or hydrogen gas because, at the same pressure, these gases are less dense than air; thus, the balloon rises to high altitudes. As altitude increases, atmospheric pressure decreases and the balloon expands. Eventually, the balloon bursts and a parachute opens to bring the radiosonde safely back to the ground. Thus, the use of weather balloons relies on a relationship between gas pressure and volume.

Activity 11.1 Cartesian Diver

When the pressure exerted on a gas increases, the volume of the gas decreases. Similarly, when the pressure exerted on a gas decreases, the volume of the gas increases. In this activity, you will construct a device called a Cartesian diver to monitor changes in the volume of air as a result of changes in external pressure on the air.

Safety Precautions

- Ensure that the bottle cap is secured tightly before performing this activity.

Materials

- water
- 750 mL or 2 L plastic bottle with cap
- eyedropper

Procedure

1. Fill the bottle to the top with water.
2. Fill the eye dropper approximately half-full with water, and place it in the bottle with its open end down. The eye dropper should float with one end just barely above the surface of the water. If necessary, add water to the eyedropper until it is barely floating. Fasten the lid of the bottle tightly.
3. Squeeze the bottle gently. Observe the eyedropper and its contents. Experiment with the device by varying the degree of compression on the bottle.

Questions

1. When you squeezed the bottle, what happened to the contents of the eyedropper? What happened to the position of the eye dropper? Provide an explanation for the motion you observed, based on the relationship between pressure and volume of a gas.
2. When you released the bottle, what happened to the contents of the eyedropper? What happened to the position of the eyedropper? Provide an explanation for the motion you observed, based on the relationship between pressure and volume of a gas.
3. How might the results of this activity relate to the ability of submarines to surface and dive?

Observations Leading to Boyle's Law

How can meteorologists predict the altitude at which a weather balloon will burst? If the balloon is designed to burst when it reaches three times the volume to which it was inflated before its release, how can they know the altitude at which that will occur? Decisions related to these questions are based on studies of the relationship between the volume of a gas and its pressure that were first published by Irish scientist Robert Boyle in 1662.

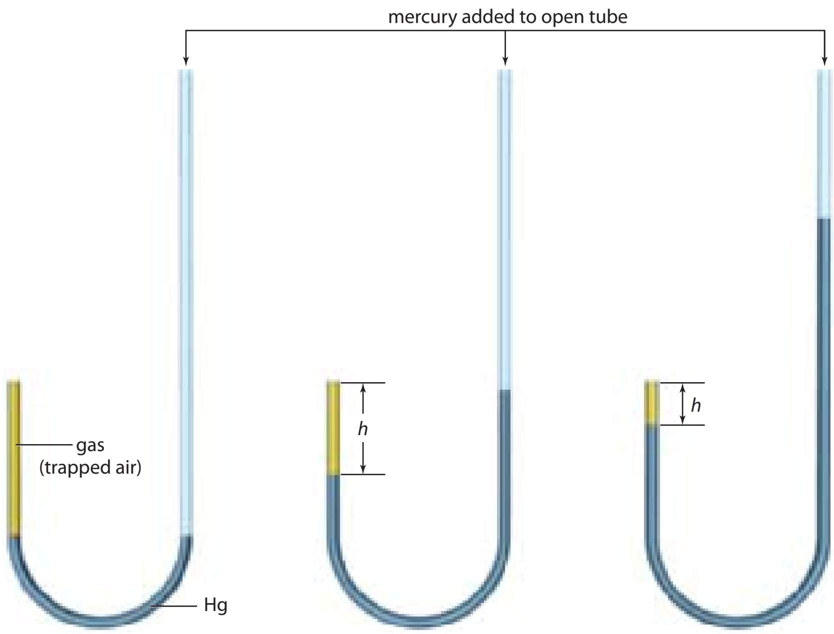

Figure 11.10 In Boyle's experiment, the left side of the U-tube is sealed and mercury has trapped air in the tube. As more mercury is added to the right side of the tube, the mercury increases the pressure on the trapped gas (air), causing its volume to decrease, demonstrated by the decreased height, *h*, of the column of trapped air in the tube..

Robert Boyle (1627–1691) studied the relationship between the pressure and volume of a gas, while the amount of the gas and the temperature of the gas were kept constant. By making careful measurements of the volume of a trapped gas, he described what happened when the pressure exerted on the gas was increased. **Figure 11.10** shows an apparatus like the one Boyle used. He measured the height of the column of trapped gas and the height of the column of mercury. The height of the mercury column is directly related to the pressure it exerts on the trapped gas. Therefore, Boyle was able to infer the relationship between the pressure on the air and its volume.

Boyle showed that if the temperature and the amount of gas are constant, an increase in external pressure on a gas causes the volume of the gas to decrease by the same factor. For example, at constant temperature, if the pressure of a gas doubles, the volume of the gas decreases by one-half. Similarly, if the pressure on a gas is reduced by one-half, the volume of the gas doubles. These observations led to **Boyle's law**, which states that the volume of a fixed amount of gas at a constant temperature varies inversely with the pressure.

$$V \propto \frac{1}{P}$$

Suggested Investigation

Inquiry Investigation 11-A, Studying Boyle's Law

Boyle's law a gas law stating that the volume of a fixed amount of gas at a constant temperature is inversely proportional to the applied (external) pressure on the gas: $V \propto \frac{1}{P}$

Developing a Mathematical Expression of Boyle's Law

A relationship in which one variable increases proportionally as the other variable decreases is known mathematically as an inverse proportion. As Boyle observed, in the case of the volume and pressure of a gas, volume, V, decreases as pressure, P, increases. Thus, the relationship can be expressed as $V \propto \frac{1}{P}$, where α is the symbol for proportionality. The graphs in **Figure 11.11** help illustrate how the mathematical equation for Boyle's law is developed. Graph A represents volume, V, versus pressure, P, and the graph's shape is typical of an inversely proportional relationship. If the relationship is in fact an inverse proportion, you should get a straight line by plotting volume against the inverse of pressure, $\frac{1}{P}$, which is evident in Graph B.

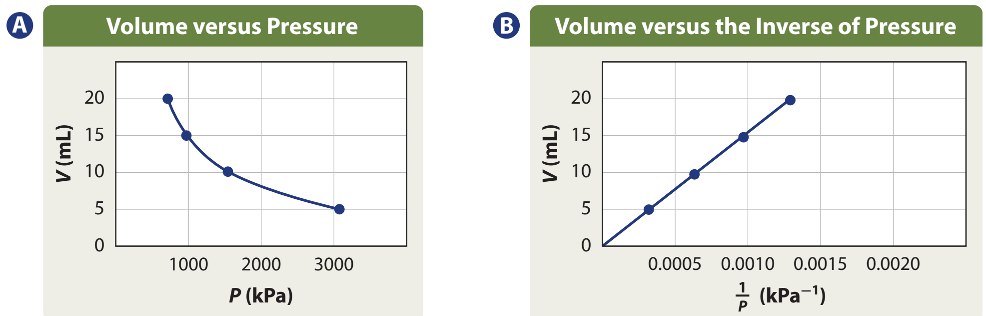

Figure 11.11 The graph for volume versus pressure **(A)** shows an inverse relationship. When you plot volume versus the inverse of pressure **(B)**, you get a straight line.

You can use Graph B in **Figure 11.11** to write a linear relationship relating volume and pressure of a gas. From your study of mathematics, you know that the general expression for a straight line is $y = mx + b$, where m is the slope of the line and b is the y-intercept. This information allows you to develop Boyle's law in mathematical form by following a few steps, as shown below.

Begin with the general expression for a straight line.	$y = mx + b$
In Graph B, the y-axis represents volume, V, and the x-axis represents the inverse of pressure, $\frac{1}{P}$. Use these values to rewrite the expression.	$V = m\left(\frac{1}{P}\right) + b$
The symbol m represents the slope of the line and b is the y-intercept. From the graph, you can see that the line passes through the origin. Thus, $b = 0$.	$V = m\left(\frac{1}{P}\right)$
Multiply both sides of the equation by P. This shows that PV is equal to a constant, which is the slope of the line.	$PV = \cancel{P}m\left(\frac{1}{\cancel{P}}\right)$ $PV = m$
Let P_1V_1 represent pressure and volume at one data point on the graph, and let P_2V_2 represent pressure and volume at a second data point. The product of pressure and volume at each point equals the constant, m.	$P_1V_1 = m$ and $P_2V_2 = m$
Because the products of P_1V_1 and P_2V_2 are equal to the same constant, they are equal to each other.	$P_1V_1 = P_2V_2$

Therefore, the mathematical expression for Boyle's law is

$$P_1V_1 = P_2V_2$$

Remember, this mathematical relationship only applies if the amount of gas and the temperature remain constant. You will use this expression of Boyle's law to understand other gas laws that you will explore in the rest of this chapter and in the next chapter.

The following Sample Problem and Practice Problems will reinforce your understanding of Boyle's law.

Sample Problem

Using Boyle's Law to Calculate Volume

Problem

A weather balloon with a volume of 2.00×10^3 L at a pressure of 96.3 kPa rises to an altitude of 1.00×10^3 m, where the atmospheric pressure is measured to be 60.8 kPa. Assuming there is no change in temperature or amount of gas, what is the final volume of the weather balloon?

What Is Required?

You need to find the volume, V_2, after the pressure on the balloon has decreased.

What Is Given?

You know the pressure and volume for the first set of conditions and the pressure for the final set of conditions.

$P_1 = 96.3$ kPa
$V_1 = 2.00 \times 10^3$
$P_2 = 60.8$ kPa

You know the temperature does not change.

Plan Your Strategy	Act on Your Strategy
Pressure and volume are changing, at constant temperature and amount of gas. Therefore, use the equation for Boyle's law.	$P_1V_1 = P_2V_2$
Isolate the variable V_2 by dividing each side of the equation by P_2.	$P_1V_1 = P_2V_2$ $\frac{P_1V_1}{P_2} = \frac{\cancel{P_2}V_2}{\cancel{P_2}}$ $\frac{P_1V_1}{P_2} = V_2$
Substitute numbers and units for the known variables in the formula and solve. Make certain that the same units for pressure are used in the equation.	$V_2 = \frac{P_1V_1}{P_2}$ $= \frac{(96.3\ \cancel{\text{kPa}})(2.00 \times 10^3\ \text{L})}{60.8\ \cancel{\text{kPa}}}$ $= 3.17 \times 10^3$ L

According to Boyle's law, when the amount and temperature of a gas are constant, there is an inverse relationship between the pressure and volume of a gas: $V \propto \frac{1}{P}$

Alternative Solution

Plan Your Strategy	Act on Your Strategy
According to Boyle's law, a decrease in pressure will cause an increase in volume. Determine the ratio of the initial pressure and the final pressure that is greater than 1.	$P_1 = 96.3$ kPa $P_2 = 60.8$ kPa pressure ratio > 1 is $\frac{96.3\ \text{kPa}}{60.8\ \text{kPa}}$
To find the final volume, multiply the initial volume of the balloon by the ratio of the two pressures that is greater than 1.	$V_2 = V_1 \times$ pressure ratio $= (2.00 \times 10^3\ \text{L}) \times \frac{96.3\ \cancel{\text{kPa}}}{60.8\ \cancel{\text{kPa}}}$ $= 3.17 \times 10^3$ L

Check Your Solution

The units cancel out to leave the correct unit of volume, L. You would expect the volume to increase when the pressure decreases, which is represented by the value determined.

Practice Problems

Note: Assume that the temperature and amount of gas are constant in all of the following problems.

1. 1.00 L of a gas at 1.00 atm pressure is compressed to 0.437 L. What is the new pressure of the gas?
2. A container with a volume of 60.0 mL holds a sample of gas. The gas is at a pressure of 99.5 kPa. If the container is compressed to one-quarter of its volume, what is the pressure of the gas in the container?
3. Atmospheric pressure on the peak of Mount Everest can be as low as 0.20 atm. If the volume of an oxygen tank is 10.0 L, at what pressure must the tank be filled so that the gas inside would occupy a volume of 1.2×10^3 L at this pressure?
4. If a person has 2.0×10^2 mL of trapped intestinal gas at an atmospheric pressure of 0.98 atm, what would the volume of gas be (in litres) at a higher altitude that has an atmospheric pressure of 0.72 atm?
5. Decaying vegetation at the bottom of a pond contains trapped methane gas. 5.5×10^2 mL of gas are released. When the gas rises to the surface, it now occupies 7.0×10^2 mL. If the surface pressure is 101 kPa, what was the pressure at the bottom of the pond?
6. The volume of carbon dioxide in a fire extinguisher is 25.5 L. The pressure of the gas in this can is 260 psi. What is the volume of carbon dioxide released when sprayed if the room pressure is 15 psi?
7. A 50.0 mL sample of hydrogen gas is collected at standard atmospheric pressure. What is the volume of the gas if it is compressed to a pressure of 3.50 atm?
8. A portable air compressor has an air capacity of 15.2 L and an interior pressure of 110 psi. If all the air in the tank is released, what volume will that air occupy at an atmospheric pressure of 102 kPa?
9. A scuba tank with a volume of 10.0 L holds air at a pressure of 1.75×10^4 kPa. What volume of air at an atmospheric pressure of 101 kPa was compressed into the tank if the temperature of the air in the tank is the same as the temperature of the air before it was compressed?
10. An oxygen tank has a volume of 45 L and is pressurized to 1200 psi.
    a. What volume of gas would be released at 765 torr?
    b. If the flow of gas from the tank is 6.5 L per minute, how long would the tank last?

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Kinetic Molecular Theory and Boyle's Law

Pressure on the walls of a gas-filled container is caused by collisions of gas molecules with the walls. Each collision of a gas molecule exerts a force on the wall. The average force exerted by all the gas molecules divided by the surface area of the container is equivalent to the pressure on the walls of the container. Examine **Figure 11.12** to see what happens when you change the external pressure on the gas. The containers have pistons that will move until the external pressure and the internal pressure are equal. If you increase the external pressure, the piston will move down, reducing the volume available to the gas molecules. The gas molecules are now closer together and collide with one another and the walls of the container more often. As the number of collisions over time increases, the average force exerted by all the molecules increases; thus, the gas pressure increases. If the temperature remains constant and no gas escapes or enters, the decrease in the volume of the container will be inversely proportional to the increase in the gas pressure.

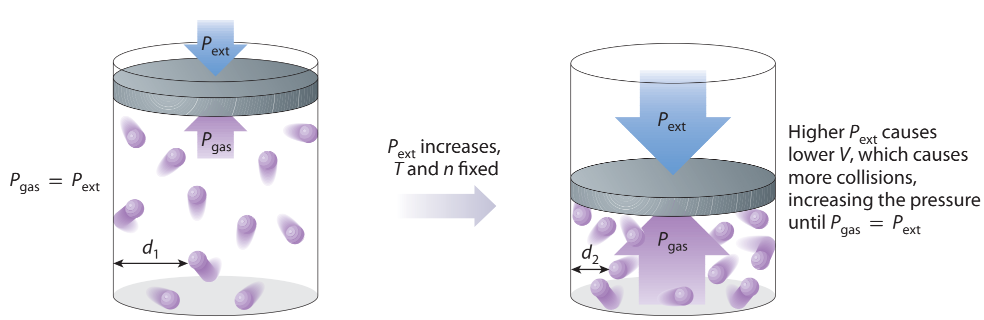


Figure 11.12 The kinetic molecular theory can explain the relationship between pressure and volume. (d_1 and d_2 represent average distances of molecules from the container wall.)

Section 11.2 REVIEW

Section Summary

- Atmospheric pressure is the force that a column of air exerts on Earth's surface divided by the area of Earth's surface at the base of that column.
- Atmospheric pressure decreases as altitude increases.
- Boyle's law states that the volume of gas is inversely proportional to the external pressure exerted on the gas when the temperature and amount of gas are constant. The equation for Boyle's law is $P_1V_1 = P_2V_2$.
- The relationship between pressure and volume of a gas can be explained using the kinetic molecular theory. As the external pressure on a gas increases, the volume of the gas decreases. As the volume decreases, the gas molecules become closer together, causing the frequency of collision of the molecules to increase, thus increasing the gas pressure.

Review Questions

1. K/U Describe how Torricelli's studies demonstrated the existence of atmospheric pressure. Include an explanation of the apparatus he used in his investigations.

2. K/U What is the relationship between atmospheric pressure and altitude?

3. A Fluid moves up a drinking straw, against gravity.

a. Explain how fluid from a glass can rise in the straw, without anyone applying suction to drink from it.

b. Would the use of a drinking straw be affected by changes in altitude? Explain.

4. K/U Explain why people who climb high mountains commonly carry bottled oxygen with them.

5. T/I For each of the following, determine which measurement is the higher pressure.

a. 1.25 atm or 101.325 kPa

b. 1.5 bar or 740 mmHg

c. 1 bar or 105 kPa

d. 800 mmHg or 1.25 atm

6. T/I A student collects hydrogen gas in a balloon fitted over the top of an Erlenmeyer flask. She records the atmospheric pressure as 98.5 kPa. Later she notices that the volume of the balloon has noticeably decreased. Initially, she hypothesizes that some of the hydrogen gas has escaped from the balloon. What data would you advise her to collect in order to confirm or refute her hypothesis? Explain your reasoning.

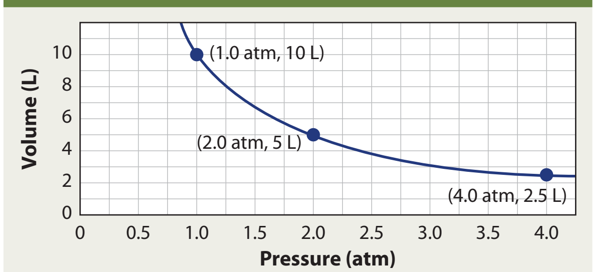
7. A One popular demonstration of gas behaviour involves putting a marshmallow in a flask and then reducing the air pressure in the flask. The marshmallow quickly swells up. How can you explain this observation?

8. A When scuba divers are rising after a dive, why is it important that they do not hold their breath?

9. K/U Describe how Robert Boyle investigated the relationship between the pressure and volume of a gas.

10. C A student performs an investigation to study the relationship between the pressure and volume of a gas. The experimental data are represented in the graph.

Pressure–Volume Changes



a. Do the data support Boyle's law? Explain.

b. How could the data be plotted so that the graph is a straight line? Draw such a graph, including *x*- and *y*-axis labels.

11. T/I The volume of a gas at 75 kPa is 4.0 L. What is the volume of the gas if the pressure increases to 100.0 kPa? Assume the temperature and amount of gas are constant.

12. T/I An air compressor tank has a volume of 60.0 L. The air from this compressor was released and found to be 2.50×10^2 L at a room air pressure of 100.0 kPa. What was the pressure of the air in the compressor tank? Assume the temperature and amount of gas are constant.

13. T/I A 20 L tank is filled with helium gas at a pressure of 10 000 kPa. How many balloons, each with a volume of 2.0 L, can be filled to a pressure of 100 kPa? Assume the temperature and amount of gas are constant.

14. C Predict what would occur if a 3 L helium balloon was taken underwater to a depth of 30 m where the pressure is 3 atm. Using kinetic molecular theory and a series of diagrams, explain any changes in volume. (Assume there is no loss of helium and the water does not change temperature.)