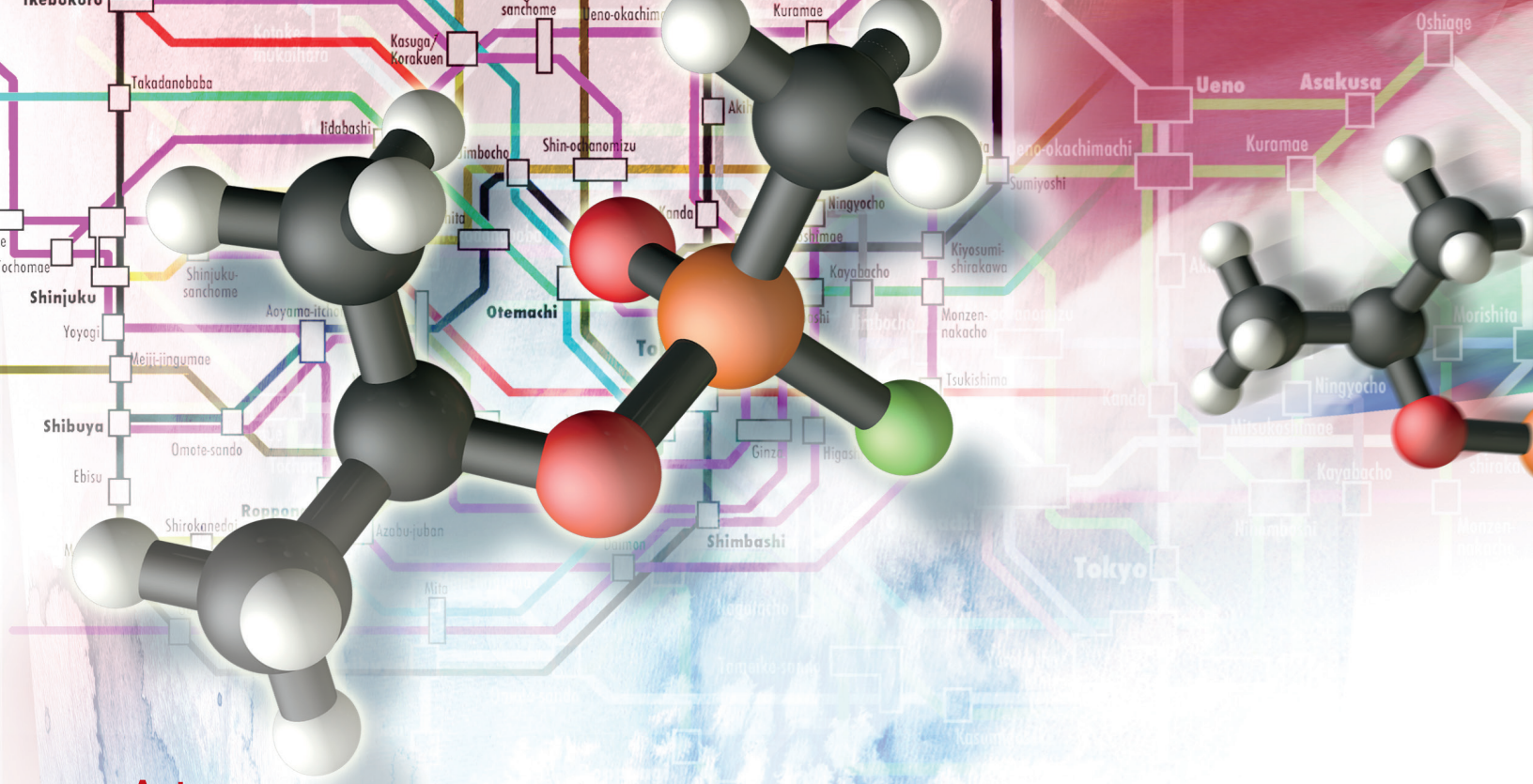
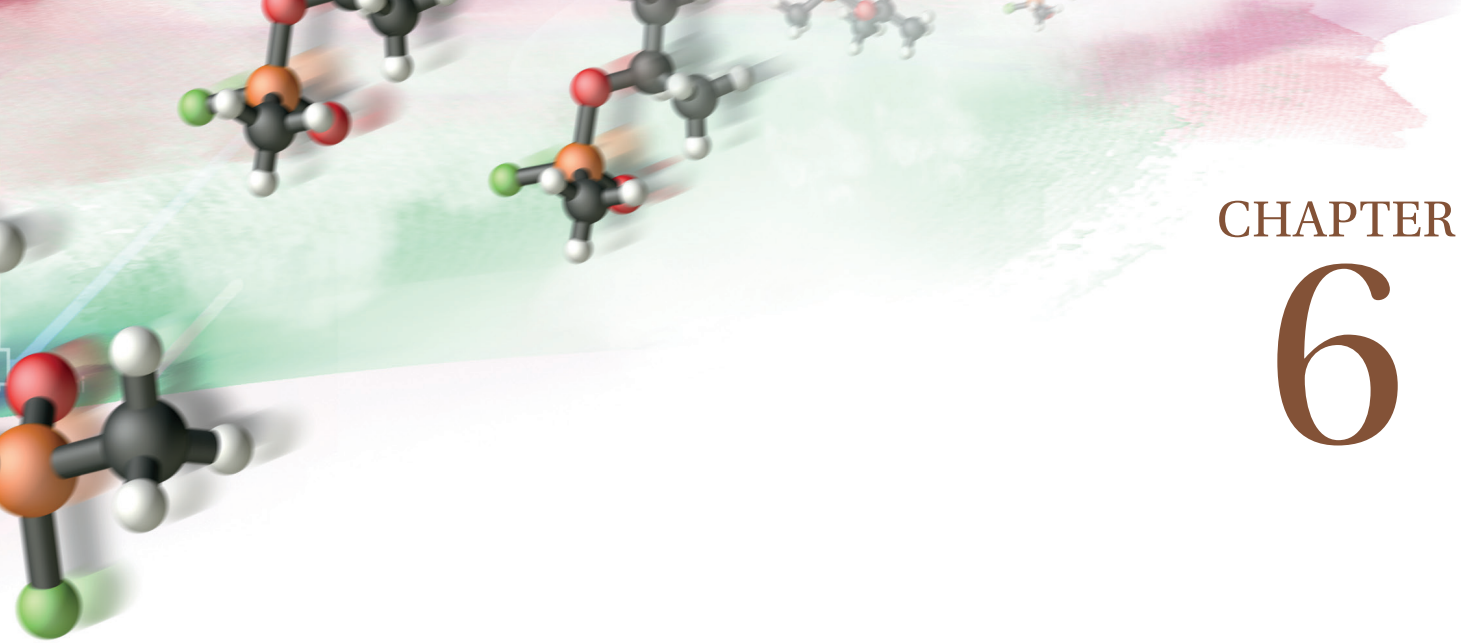


Air

A Study of the Gases in Our Atmosphere



CHAPTER

6

March 20, 1995, promised to be a typical spring day in Tokyo as people left their homes to go to work. What they did not know was that members of the Aum Shinrikyo terrorist cult were en route to various points along the Tokyo subway system. At the height of the morning rush hour, members of this group punctured containers of sarin, a deadly nerve agent that has no color or odor. The gas spread through the air in several crowded subway cars.

When sarin changes from a liquid to a gas, which it does very readily, the gas can be absorbed through the skin or the lungs and enter a person's bloodstream. Sarin quickly disrupts the normal functioning of the nervous system, causing paralysis, convulsions, loss of vision, and violent headaches. In doses greater than 0.5 milligrams, exposure is usually fatal. During the Tokyo subway attack,

many of the people who were exposed to near-fatal levels of sarin escaped death but suffered permanent neurological damage. Those exposed to smaller quantities of sarin recovered completely. All told, the attack resulted in the death of 12 people; more than 5000 others were injured. The fate of any one person riding the Tokyo subway on that horrifying day depended on the behavior of the gases around that person.

AP Images/Ju wen/Imaginechina

Our bodies are surrounded by gases at all times. Clean, dry air is colorless and odorless and contains about 21% oxygen, 78% nitrogen, and small amounts of other benign gaseous atoms and molecules (**Table 6.1**). In the same way that we would not eat contaminated food or drink contaminated water, we don't want to see or smell the air that we breathe. Gases we can see or smell may be cause for alarm. Gases we can see may be laden with undesirable specks of airborne pollution (see the photo of people cycling through polluted air in Tiananmen Square in Beijing, China). Gases we can smell contain molecules that trigger the olfactory nerve. This nerve transmits odor information, which may include signals of possible contamination, from the nose to the brain.

In this chapter we explore gases, such as air, the gas that constantly surrounds us. To learn about air, though, we must first understand how gases behave. We'll do this by learning some of the rules we expect gases to follow.

Table 6.1 Composition of Clean Air at Sea Level

Atoms or Molecules	Percentage (%)
Nitrogen molecules, N_2	78.1
Oxygen molecules, O_2	20.9
Argon atoms, Ar	0.934
Carbon dioxide molecules, CO_2	0.036
Trace gases	<0.03

6.1 The Nature of Gases

Of the three phases of matter, gases are the simplest.

Of the three phases of matter, gases may be the simplest. They are simple because, unlike liquids and solids, they tend to be completely homogeneous: except under special circumstances, gases mix thoroughly with one another. For example, although all the gases listed in Table 6.1 are present in clean air, clean air is completely homogeneous because it's the nature of all gases to blend together.

flashback
Recall from Section 3.5 that there are three phases of matter: gas, liquid, and solid.

Gases are also simple because individual atoms or molecules within a gas typically do not interact with one another. Sure, they collide with one another now and then, but for the purposes of our discussion in this chapter, we can assume that the atoms or molecules of a gas do not change when they collide. In other words, they do not undergo any chemical reactions. This means it's safe to assume that all of the atoms or molecules in a gas act the same way, no matter what elements they contain. For this reason, we can think of any of the atoms or molecules of a gas as a generic "particle of a gas." These gas particles behave like tiny billiard balls that collide and bounce.

Although they are homogeneous, and particles of a gas do not interact with one another, gases are high maintenance. **Figure 6.1** highlights the differences between the three phases of matter. One important difference between liquids, solids, and

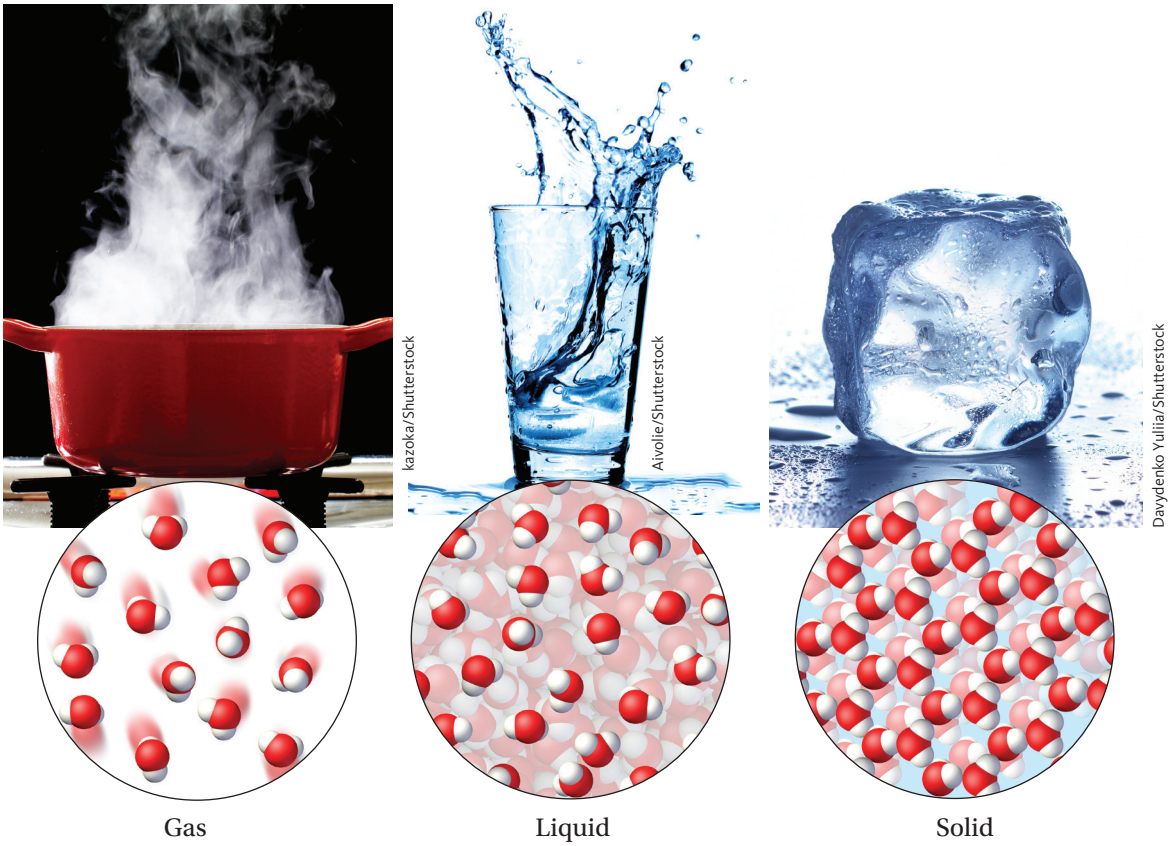


Figure 6.1 The Three Phases of Matter

(a) In a gas, the particles are far apart from one another and fill all the available space in the container holding them. (b) In a liquid, molecules are close together but move relative to one another. The liquid conforms to the shape of the container but does not fill the entire container. (c) In a solid, molecules are tightly packed in a regular array and move very little relative to each other. The solid retains its own shape regardless of the shape of the container holding it.

gases is how they must be stored. We can leave a sample of most solids and liquids sitting out in the open and not worry about losing any of it (except small amounts that are lost when the liquid evaporates or a solid sublimes—that is, turns from a solid directly into a gas). That's because, within solids and liquids, there are forces that hold together individual units. With gases, though, the forces of attraction are not strong enough to keep the individual gas molecules together. That is why we store most gases in airtight containers to prevent leakage.

Think about gas-filled things we see or use every day: a tire, a pipe carrying natural gas into a building, a neon sign, an unopened can of soda. In all these cases, the container must be leakproof. If a leak develops, the product or device ceases to do its job or ceases to be useful. Sometimes this is no big deal, as with a burst soda can. But when the gas in the container has the potential for harm, as in the case of highly flammable natural gas, escaping gas can lead to an explosion.

QUESTION 6.1 Imagine you have a glass containing ice cubes and some water. If the glass is tightly covered with a lid, how many phases of water exist in the glass?

ANSWER 6.1 All three phases. Liquid water is mixed with ice (solid water), and water vapor (gaseous water) exists in the space above the liquid surface.

For more practice with problems like this one, check out end-of-chapter Question 3.

FOLLOW-UP QUESTION 6.1 Identify the phase of matter described in each example. (a) the phase that conforms to the shape of its container and has a surface (b) the phase that is always homogenous (c) the phase that is least likely to disappear from an open container if left undisturbed.

Gas particles move fast and are far apart from one another.

All gas atoms and molecules move very fast. A typical nitrogen molecule at room temperature, for instance, moves at the supersonic speed of about 1850 kilometers per hour (km/h) or 514 meters per second (about 1150 miles per hour), nearly as fast as the fastest jet airplanes. If we could point a radar gun at the individual molecules in a sample of pure nitrogen, however, we would find that they are not all moving at the same speed. In fact, if we could measure the speed of each individual molecule and keep a tally of the number of molecules moving at each speed, we would end up with a curve. The average speed of all the molecules would be somewhere near the middle of the curve, as shown in **Figure 6.2**. Most of the molecules in the sample are moving at or near the average speed, but there are outliers that move either much faster or much slower.

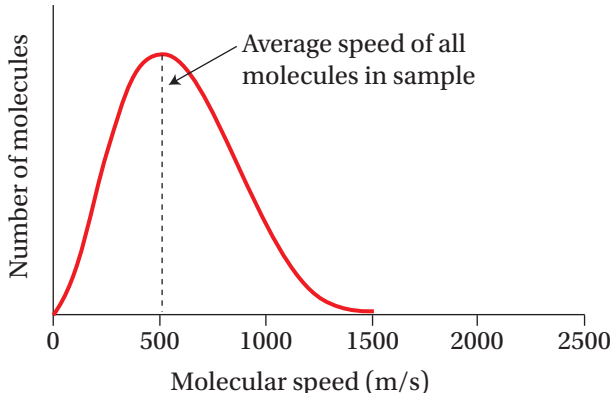


Figure 6.2 The Molecules in a Gas Sample Move at Various Speeds

This graph shows the speed of all the molecules in a sample of nitrogen gas at 27°C. The speeds are distributed in a curve because some molecules move at speeds lower than the average speed of all the molecules in the sample while other molecules move at speeds higher than the average speed.

Gas molecules are also very far apart and take up only a small percentage of the total space of the container holding them. Even so, it's inevitable for them to come into contact with one another occasionally. The collisions that occur in a gas are assumed to be completely elastic: The molecules simply bounce off each other. We can think of the mole-

cules in a gas as tiny, high-speed billiard balls that are in continuous, random motion through space.

The random motion of gas molecules leads to some interesting effects. For instance, although it's easy to see your brother enter the house after cleaning fish in the backyard, it takes a while before you smell him, because the smelly gas particles he picked up from the fish do not travel directly from his body to your nose. Rather, they take a zigzag path across the room, all the while colliding with other gas particles as well as walls and tables and lamps (**Figure 6.3**).

double_p/iStock/Getty Images

Figure 6.3 A Real Stinker

The molecules associated with strong odors eventually make their way to your nose.

Fast-moving gas particles mix together quickly and completely.

Gas molecules move into every space available and completely fill any container they occupy. Imagine we introduce a new molecule to a population of gas molecules speeding around in a container. The new molecule swiftly becomes integrated and assimilated into the existing mixture of fast-moving molecules. This process, known as **diffusion,** is analogous to a busy highway where cars entering the highway quickly reach the same speed as the other moving cars. Once a new car enters the highway, it becomes part of the traffic flow, where all the cars are moving at slightly different individual speeds but at a similar average speed.

The assimilation of one gas into another has important practical consequences because the air we breathe naturally mixes with other gases quickly and completely. As described at the start of this chapter, the canisters of sarin used in the Tokyo subway released poisonous gas into the air. Just like cars merging into fast-moving traffic, the deadly molecules became mixed in with the rest of the molecules in the air breathed by thousands of commuters. Evidence gathered after the attack showed that people closest to the canisters had the most severe symptoms of nerve gas poisoning. This finding is reasonable because, as the poison gas molecules moved farther from a canister, they became farther apart from one another as they diffused into the air and became diluted. A passenger farther from a canister therefore inhaled air containing fewer sarin molecules, and a passenger closer to the canister inhaled more. The closer passenger would reach the fatal limit of sarin with fewer breaths.

QUESTION 6.2 Which of these statements about gases is untrue? (a) Gas particles move fast. (b) In a sample of a gas, all gas particles move at the same speed. (c) Gas particles mix completely with each other. (d) Gas particles bounce off each other and off container walls.

ANSWER 6.2 (b) In a sample of gas, there is a distribution of speeds for gas particles, as shown in Figure 6.2.

For more practice with problems like this one, check out end-of-chapter Question 2.

FOLLOW-UP QUESTION 6.2 True or false? When gases diffuse, gas molecules do not collide with walls or with other gas particles.

6.2 Pressure

The pressure that a gas exerts is related to the collisions the gas particles make with their container.

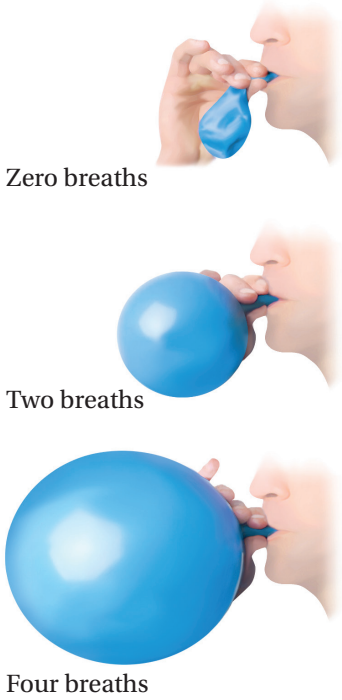


Figure 6.4 Three Balloons

Consider a simple gas-holding device: a balloon. Suppose that we begin with three identical, uninflated balloons. If we take the first uninflated balloon and tie off the end of it, the balloon contains only the air that was in the room before we tied it. Next, we take the second uninflated balloon and blow two deep breaths into it and tie it off. Finally, we take the third uninflated balloon and blow into it until we think it might pop. Imagine that this process takes four breaths. Our three balloons look very different now: one is flaccid, one is partly filled, and the third is completely filled, as shown in **Figure 6.4**.

What makes these three balloons different? They are different because they all contain different amounts of gas. (Remember, we can ignore the fact that the gas we blew into the balloon is a different mixture of gases than the gases in the room. We consider all gas particles to be equal.) The first balloon contains the fewest gas particles, the third balloon contains the most gas particles, and the second balloon contains some intermediate number of particles. So how do gas particles cause a balloon to inflate?

Inside the second balloon, which is now partially inflated, the gas particles that we forced in with the strength of our lungs are moving very fast, as gas particles tend to do. But these gas particles are not free to roam anywhere; they encounter the balloon's inside surfaces constantly. And every time a particle of gas hits the balloon's inside surface, it pushes it outward very slightly and then bounces away to the inside of the balloon. Quite soon, each particle is colliding with the surface again, then again, and again. Billions and billions of particles inside the balloon are behaving like this, and they're all pushing against the inside surfaces of the balloon over and over again. This is what inflates the second balloon and gives it some shape.

The third balloon is a more extreme example of this phenomenon. It contains even more particles, and all of those particles, together, are pushing even harder on the inside surfaces of the balloon. Thus, the third balloon is fully inflated, as illustrated in **Figure 6.5**. Eventually, if we blow enough gas particles into a balloon, the force of the particles pushing on the interior of the balloon is too much. The balloon material is not strong enough to withstand the force, and it pops!

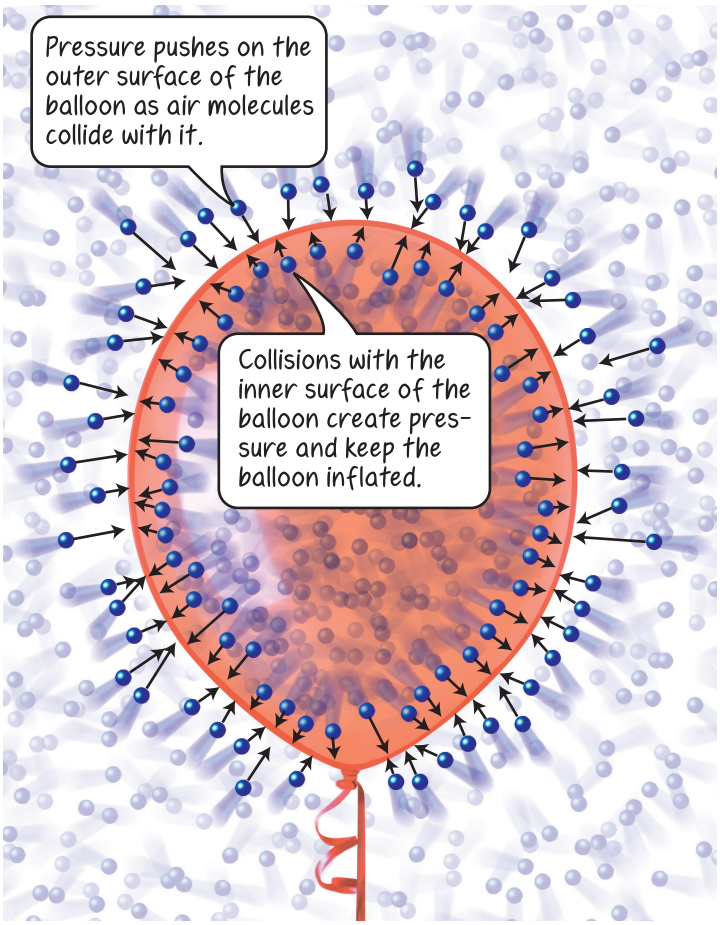


Figure 6.5 Why a Balloon Inflates

This balloon is inflated because the gas particles within it push against the inside walls whenever particles collide with them. The pressure inside the balloon pushes against the pressure of the atmosphere on the outside of the balloon.

Figure 6.6 An Uninflated Tire versus an Inflated Tire

(a) A flat tire contains fewer gas particles than a full tire. As a result, less pressure is pushing against the inside tire walls. (b) A fully inflated tire contains enough gas particles to press on the inside tire walls and keep the tire inflated.

Pressure is a force applied to a surface.

Now we can see why a balloon that contains more gas particles inflates more. The more gas particles, the more collisions on its inside surfaces, the more inflated the balloon must be. We refer to the constant pushing by gas particles that's taking place on the interior surface of the balloon as pressure. **Pressure** is a force, such as the force from gas particles, applied to a surface, such as the inside surface of a balloon.

The term *pressure* is not used exclusively for gas particles that are pushing on a balloon's inside surface. We can use it for a hand pressing on the surface of a door, or a foot stepping on an accelerator pedal, or a thumb pushing onto a fingerprint card. For this reason, we encounter units of pressure like pounds per square inch, also known as psi. This pressure unit, which we might see on a tire gauge, indicates the number of pounds of force being pressed into an area—in this case, each square inch of the surface. **Figure 6.6** shows how the pressure of a tire changes as it is filled with air.

Atmospheric pressure changes with altitude.

Imagine you live at sea level—say, in San Francisco. After deciding to take a skiing vacation in the Colorado Rockies, you pack your bags and hop on a plane. When you arrive in Colorado and unpack your bags, you notice that your plastic shampoo

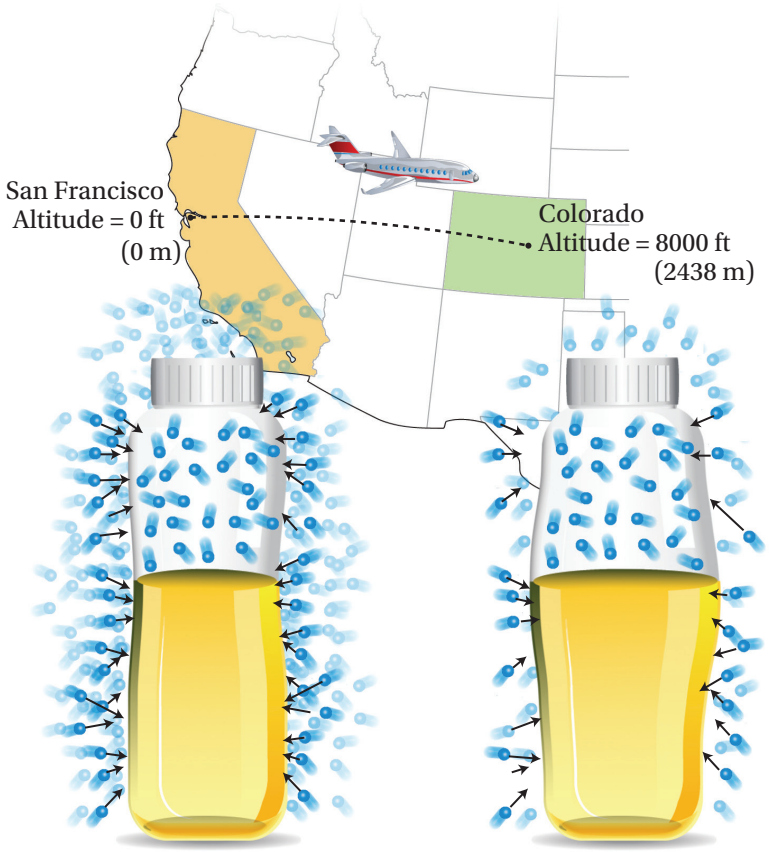


Figure 6.7 I Left My Air Pressure in San Francisco

When the cap is put on a plastic shampoo bottle at sea level, the air pressure inside the bottle is the same as the air pressure outside. Later, at an altitude where the air pressure is lower than at sea level, the air trapped inside the bottle exerts enough pressure on the inside walls to distort the shape of the container, which is flexible.

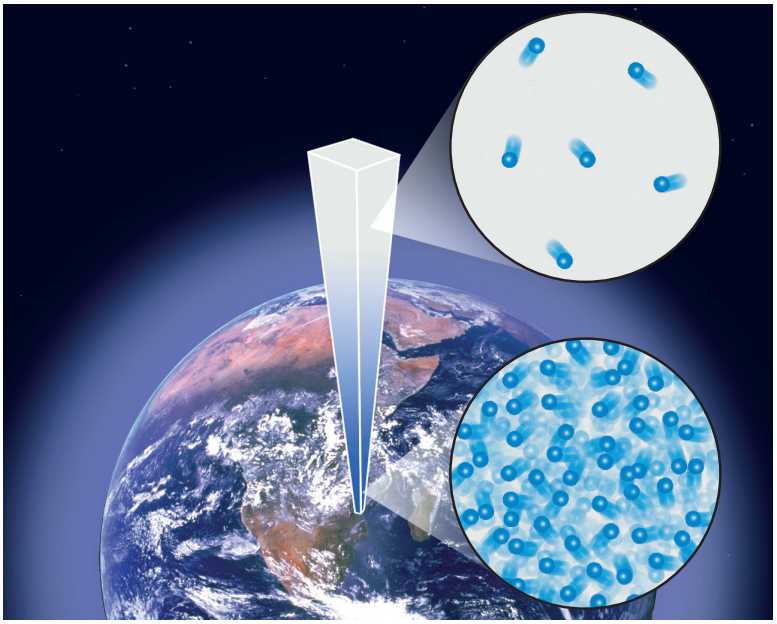


Figure 6.8 Into Thin Air

An imaginary column of air starting at Earth's surface (sea level) and extending into the atmosphere has fewer and fewer air molecules and atoms per volume as the distance from sea level increases. In other words, the air density decreases with increasing altitude.

bottle is bigger than it was when you packed it back home at sea level (**Figure 6.7**). You carefully open it, and a gentle whish of air comes out. On the return trip, you arrive in San Francisco to find that your shampoo bottle is now smaller than it was when you packed it in Colorado. What's going on?

Although we cannot see them, atoms and molecules in the air are bouncing around and ricocheting off us and off everything around us. Our bodies provide a surface for bouncing gas molecules that is no different from the interior surface of a balloon or a tire. Every surface exposed to air is constantly bombarded by the gas molecules making up the air. However, surfaces exposed to air at different altitudes are bombarded with different numbers of molecules.

To understand why, we must know two things: (1) gravity pulls everything downward toward Earth's surface; and (2) atoms and molecules are not exempt from the pull of gravity. The atoms and molecules in air are pulled downward just like everything else. So we can imagine that, at sea level, the molecules in air are relatively close together because a given volume contains more of them than that same volume does in, say, the mountains of Colorado. Put another way, air is denser at sea level than at high altitude, and its density changes through all points in between (**Figure 6.8**).

We can solve the mystery of the distorted shampoo bottle by considering the density of air at the two altitudes. The denser air in San Francisco means that, relative to Colorado, more collisions are occurring between air molecules and the outside of the bottle. This means that the pressure on the outside of the bottle is greater in San Francisco than in Colorado. When you pack your shampoo in San Francisco, the density of air is the same on the inside and outside of the bottle, and so the bottle maintains its normal shape. When you carry that sealed bottle to the mountains, though, the pressure on the inside of the bottle (the sea level pressure) is now greater than

the pressure on the outside (the high-altitude pressure). Molecules in the air inside the bottle push outward, and the bottle expands.

QUESTION 6.3 Imagine you arrive in Colorado and your shampoo bottle has leaked all over the new ski clothes in your suitcase. Describe what happened using the concepts discussed in this section of the chapter.

ANSWER 6.3 When you get to Colorado, the pressure inside the bottle is higher than the pressure outside. The inside pressure pushes against the shampoo, forcing it up against the covered hole in the bottle cap. The pressure pushing on the cap from the outside of the bottle is lower than the pressure pushing from the inside of the bottle. Thanks to this difference in pressure, the cap is pushed outward and pops off.

For more practice with problems like this one, check out end-of-chapter Question 15.

FOLLOW-UP QUESTION 6.3 On average, are the particles of gas in the air closer together (denser) or farther apart (less dense) on top of Mount Whitney, the tallest mountain in California, than they are at the beach in San Diego?

A gas particle's mean free path is the distance it travels between collisions.

One way to imagine the density of gas is to think about the motion of a handful of gas particles. We know that they travel at very high speeds and randomly careen from one place to another, colliding occasionally with other gas particles and with the walls of their container. It follows that those collisions are more frequent in a high-density gas than in a low-density gas. We can quantify this difference in terms of the **mean free path** of a molecule, defined as the average distance the molecule travels between collisions.

For example, on average, the mean free path of the atoms and molecules in air at sea level is about 60 nanometers (nm). This means that, on average, a particle in air collides with something else every 60 nm that it travels. At high altitude, though, we would expect the mean free path to be greater because at high altitude there are fewer molecules in a given volume. And this is exactly what we find: the mean free path of air particles on the summit of Mount Everest is about 180 nm, three times longer than the mean free path at sea level.

At the top of Everest, molecules in the air are about three times farther apart than at sea level. Thus, a single breath of air on the summit of Everest, represented by the clear box on the left of **Figure 6.9,** supplies the body

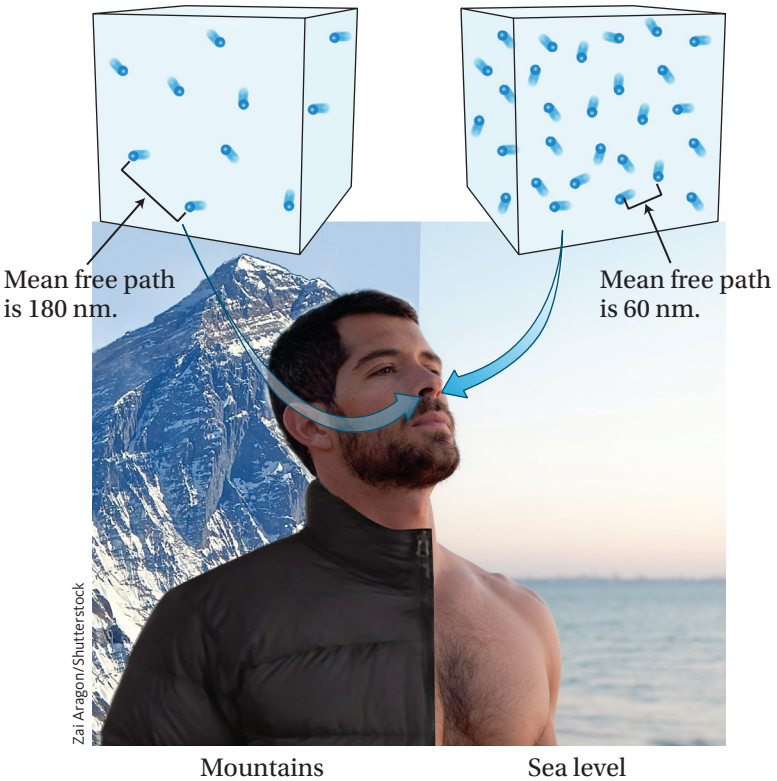


Figure 6.9 Breathing at Two Altitudes

Consider a person in the mountains (left) and at sea level (right). The clear boxes represent one big breath of air at each location, and the particles inside show the mean free path at these two elevations. Gas particles in air are farther apart in the mountains (left) compared to sea level (right).

natureBOX • Is Natural Gas the Ideal Energy Resource?

In 2005, the United States was poised to make a dramatic shift in its energy use. Many people favored ending military and political conflicts stemming from our foreign oil dependence, and natural gas entered the arena as the new, cleaner energy source.

Natural gas is a mixture of gases, but it is predominantly composed of methane. Methane is a molecule with one carbon atom surrounded by four hydrogen atoms, and it is a fossil fuel because it comes from organic matter decay in the earth. It is considered relatively clean because it is a gas and is not mixed with undesirable substances like mercury and sulfur, which make coal and crude oil "dirty."

Since 2005, natural gas use has mushroomed in the United States. In the areas that have the most abundant underground natural gas—Pennsylvania, Ohio, and West Virginia—the natural gas withdrawal has increased from about 2 billion cubic feet per day in 2009 to a whopping 24 billion cubic feet per day in October 2017.

Natural gas is most often extracted using *hydraulic fracturing*, also known as *fracking*. As the figure illustrates, a well is drilled down through the water table and into *shale*, a dense rock whose fissures hold natural gas. The drill then turns and drills horizontally into the shale. When a watery mixture is pumped into the shale at very high pressure, the rock fractures. Natural gas is sucked out and up, where it is collected at the well head.

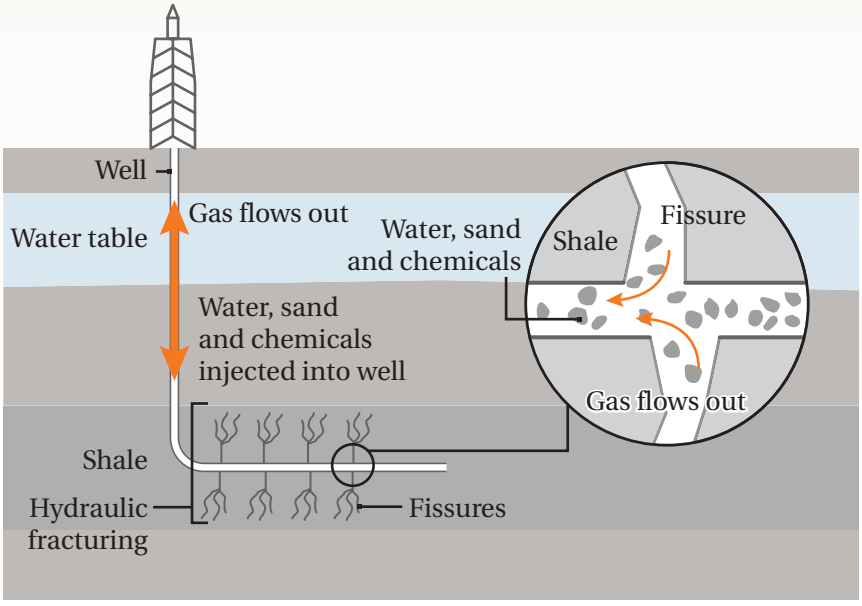
If fracking produces cleaner natural gas and reduces our foreign oil reliance, then why are some people against it? The clearest reason is because natural gas, which is mostly methane, produces carbon dioxide when burned. Carbon dioxide produced by burning fossil fuels causes global warming, and this temperature increase is rapidly changing our climate. Furthermore, methane, which leaks into the air unburned as part of fracking, contributes to global warming four times more than does carbon dioxide. Thus, while natural gas may decrease our foreign oil dependence, it does not solve the problem of global warming (see Chapter 12).

Additionally, fracking fractures rock using fluids, which may contain toxins like benzene, a potent carcinogen. These toxins have begun appearing in the water of residents living near drilling sites. Since fracking releases methane into the surrounding rock and into the water table, there are reports of "flammable water" coming from kitchen faucets. The water is flammable because it can be lit with a match.

Even more dramatic than flaming faucets is the increase in earthquakes in areas where fracking is common. For example, before fracking began in Oklahoma, the state saw about 1.5 earthquakes that measured 3.0 or more on the Richter scale, per year on average. Since fracking wells have been operating, this number has risen to more than 500 per year. Texas, Arkansas, and Colorado have seen a similar increase.

Given these facts, is fracking the ideal solution to our energy needs? Perhaps it is a reasonable first step that will free us from imported oil. However, as a fossil fuel, it is not renewable or sustainable. In Chapter 12, we will explore energy alternatives that *are* renewable and sustainable.

Shale gas extraction



with fewer particles of air than the same-sized breath at sea level, represented by the clear box on the right in Figure 6.9. Because about one in five air particles are oxygen molecules, this also means there are fewer molecules of oxygen on Everest as compared to a lower elevation. This is why most climbers who scale Everest rely on supplemental oxygen to get them to the top.

wait a minute . . .

Is the percentage of oxygen in air lower at high altitude than at sea level?

No. The composition of air is roughly the same on a mountaintop and at sea level; in both places, air contains about 21% oxygen. However, at high altitude there are fewer gas molecules per volume of air. One breath of air at 18,000 feet (about 5500 meters) will contain about half the number of oxygen molecules as the same breath taken at sea level. So, even though the composition of air does not change between sea level and a mountaintop, the density of the air *does* change and this is why we breathe less oxygen at high altitude.

Atmospheric pressure is the pressure exerted on us by air in the environment.

Although we cannot see it and we really do not feel it, we have evidence that air is always colliding with our bodies. As we have just seen, however, the pressure pushing on us varies from place to place. It also changes with the weather. A weather system may bring high pressure (when the pressure of air molecules on us is greater than normal) or low pressure (when the pressure of air molecules is lower than normal). The pressure exerted on us by the air in the environment, called **atmospheric pressure,** fluctuates constantly. As a part of weather prediction, meteorologists keep track of pressure fluctuations in the atmosphere. When the pressure changes, the weather changes.

How is atmospheric pressure measured? **Figure 6.10** (on p. 170) shows a rudimentary, yet accurate, version of a **barometer**—a device for measuring atmospheric pressure. A glass column sits in a pool of mercury that is open to the air and therefore subject to the force of air molecules in its environment. Because pressure is defined as the force exerted over a given surface area, the higher the pressure on the surface of the mercury in the open dish in Figure 6.10, the farther up the tube the mercury moves. The height of the mercury column in the tube is therefore proportional to the atmospheric pressure. A common unit for expressing pressure is **millimeters of mercury (mm Hg)**. This is a measurement of the

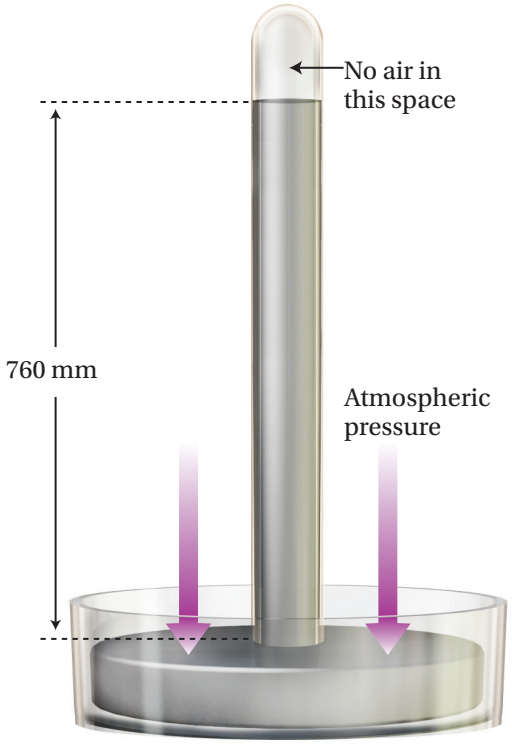


Figure 6.10 A Simple Barometer

In a barometer, atmospheric pressure pushes on the surface of the mercury liquid in the bowl, forcing the mercury up into the evacuated tube. The height of the mercury column is one way to express the magnitude of the atmospheric pressure.

distance, in millimeters, that the mercury in a barometer tube travels as a result of atmospheric pressure.

Now let's consider another unit of pressure commonly used in scientific measurements. The **atmosphere (atm)** is a pressure equal to 760 millimeters of mercury. It's a convenient unit of measure because at sea level the atmospheric pressure is typically about one atmosphere, more or less.

QUESTION 6.4 Which of these statements about pressure is untrue? (a) Air pressure does not vary with altitude. (b) Atmospheric pressure is measured with a barometer. (c) Atmospheric pressure can affect weather. (d) Pressure is a force applied to a surface.

ANSWER 6.4 (a) Air pressure *does* vary with altitude, as our shampoo bottle example demonstrates.

For more practice with problems like this one, check out end-of-chapter Question 21.

FOLLOW-UP QUESTION 6.4 Which of the following is *not* a unit of pressure discussed in this chapter? (a) millimeters of mercury (b) joules (c) psi (d) atmospheres

6.3 Variables That Affect Gases: Moles, Temperature, Volume, and Pressure

The mole allows us to count very small things, such as atoms and molecules.

Consider an imaginary box measuring 28.19 centimeters in each of its three dimensions: length, width, and height. This is approximately one foot on each side. If we do the math, we find that the volume of the box is 22.4 liters. It's also an unusual box, because it can be hermetically sealed—when it's shut, nothing can go in and nothing can come out. So, it's a perfect place to keep a gas. Now, imagine that the temperature in the box is 0°C and pressure in the box is 1.00 atmosphere. We refer to this pair of conditions as *standard temperature and pressure,* and we abbreviate it **STP**. In this case, we can say that our box is at STP. For simplicity, let's impose one final constraint on our box: the box contains a gas, and it can be any gas at all. The gas could exist in the form of molecules or just atoms, but we will think of the gas—whatever it is—as individual particles flying around inside the box.

Suppose we would like to know how many particles of gas are in the box. As it happens, under these special conditions, we know the answer to this question: there are 602,000,000,000,000,000,000,000 particles in the box. This may seem like an enormous number of particles, and it is, but it's a typical number of particles for a container like this to hold. Under reasonable conditions, in a box of a size that

would fit on your car seat, this is roughly the number of particles that are present. Now, what if we have two identical boxes rather than just one box? In that case, we have twice as many particles of gas: 1,204,000,000,000,000,000,000,000 particles in the two boxes. Imagine that we have one box again, but the new box is one half the volume of the original box. How many particles does it contain? Half as many particles: 301,000,000,000,000,000,000,000.

It's pretty straightforward to see how the number of particles changes as we change the size of the box or the number of boxes, but it can get pretty tedious to write numbers like the blue ones in the last paragraph. Luckily, we have two ways to get around this problem. First, we can use scientific notation and say that our original box contains 6.02×10^{23} particles. That is much easier! But there's an even quicker way to express this number. Rather than saying, "This box contains 6.02×10^{23} particles of gas," we can simply say, "This box contains one mole of gas."

If we have one **mole** of anything, then we have 6.02×10^{23} of those things. Now we see why our original box is special: under those exact conditions, that specific box holds *exactly* one mole of gas. For this reason, we say that the volume of that box—22.4 liters (L)—is the **molar volume** of a gas. In other words, this is the volume that holds exactly one mole of gas at STP. **Figure 6.11** shows an everyday item that has a volume of about 22.4 L.

A mole is a counting unit that makes it easy to talk about the number of particles in a quantity of a substance. It's like the word *dozen,* which is an alternate way of talking about 12 things, be they eggs or hats or snow shovels. Likewise, we can use the mole to count anything. We can have a mole of doorknobs or coffee beans or helium atoms or ballet slippers. Whatever it is, a mole is 6.02×10^{23} of that thing. The number 6.02×10^{23} is sometimes called **Avogadro's number** after Amedeo Carlo Avogadro, a nineteenth-century Italian scientist.

Although the mole can be used to count anything, it makes sense that we typically use the word *mole* when we're talking about atoms or molecules or other very tiny things. We won't often need to talk about a mole of doorknobs because, if we set out to cover Earth with a layer of doorknobs, we would need nearly 10 million Earths to distribute one whole mole of doorknobs. So, even though we *could* use moles to count doorknobs, the mole is most useful when we use it to count really tiny things, such as atoms and molecules.

flashback

Recall from Section 1.2 that scientific notation is a way to express large or small numbers using the number 10 raised to an exponent. In the number 6.02×10^{23}, the exponent is 23. We can express this number in non-exponential form by starting at the decimal point and moving it to the right 23 times. The result is 602,000,000,000,000,000,000,000.

recurring theme in chemistry 2

Things we can see with our eyes contain trillions upon trillions of atoms. Atoms are much too small to view with the naked eye.

Naypong/Shutterstock

Figure 6.11 Everyday Molar Volume

Both the items shown here have a volume of about 22.4 L. At STP, these volumes contain about one mole of gas particles.

QUESTION 6.5 Which of the following objects would likely not be counted using moles? (a) carbon atoms (b) sugar cubes (c) molecules of methane (d) electrons

ANSWER 6.5 (b) Sugar cubes are not tiny, but the rest of the items listed are. Therefore, we would probably not count sugar cubes in mole units.

For more practice with problems like this one, check out end-of-chapter Question 31.

FOLLOW-UP QUESTION 6.5 A sealed box at STP has a volume of 2240 liters which is equal to 100 times the molar volume of a gas. How many gas particles does the box contain?

Four variables dictate the behavior of a gas.

When we defined molar volume, we were careful to specify the exact conditions of the box that contained our gas. One of the conditions that we specified was temperature, because it's one of the variables that must be controlled if we want to understand how a gas is behaving. A **variable,** such as temperature, is a condition that can be changed.

How does varying temperature affect the behavior of a gas? Temperature (T) changes the speed at which gas particles move through space, and this in turn affects how quickly a given gas particle reaches the next wall it is going to collide with. This change in speed, therefore, affects the pressure of the gas. Because the pressure can change as the temperature is varied, pressure (P) is also a variable.

So far, we have seen that T and P can be altered for a sample of gas, and that changing one of these variables causes a change in the other variable. What else might change for a sample of gas? Or, put another way, what are other examples of variables? As we saw from our balloon example in Section 6.1, the balloon got bigger when we blew more gas into it. This tells us that the volume (V) the gas occupies is another variable.

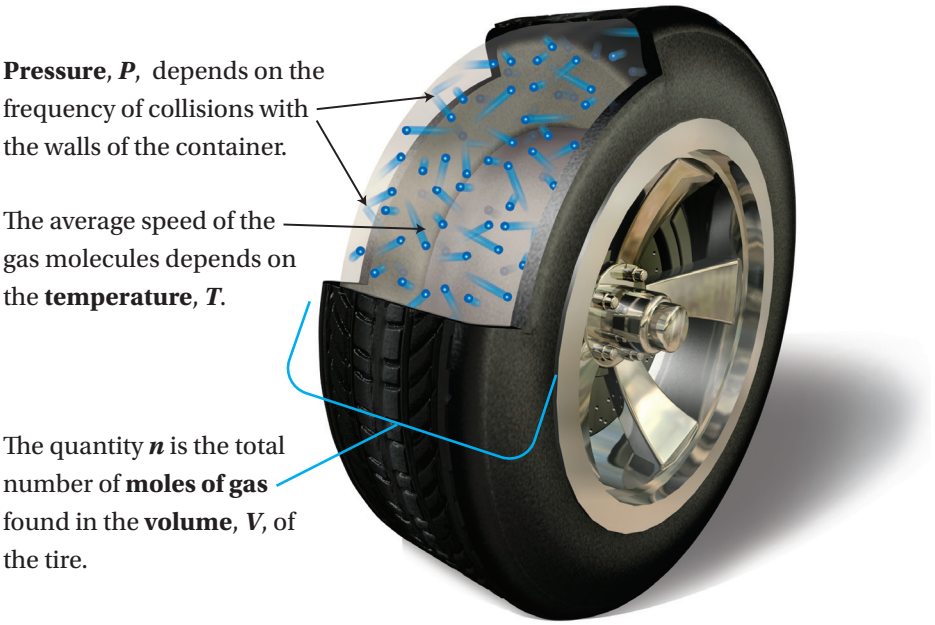


Figure 6.12 The Four Variables That Affect Gas Behavior

Pressure, temperature, volume, and number of moles all play a part in how a gas behaves.

We have just one more variable to consider. Remember that we changed the balloon's size by changing the number of gas particles in it: the more breaths we blew in, the more particles of gas were in the balloon. More gas particles mean more collisions with the inside walls of the balloon. More collisions expand the balloon. This effect tells us that the number of particles is a fourth variable that can be changed. Because the mole is the way we count particles, we use moles to keep track of particles, and we use the variable n to represent this number.

We now have a list of four variables that can change for any sample of a gas: temperature (T), pressure (P), volume (V), and the amount of gas (n), which we count in mole units. When we change any one of those variables, one or more of the other variables must also change, as we'll see in the next section. These four variables are summarized in **Figure 6.12**.

6.4 The Gas Laws: An Introduction

Pressure and volume are inversely proportional to one another.

In Section 6.3, we learned that each of our four gas variables can change as the other variables are changed. Now we will choose some specific pairs of variables, in turn, and create statements about how they change in relation to one another. We call these statements **gas laws,** but they are really nothing more than formal statements of commonsense observations about the way gases behave. You will read about several gas laws in this chapter, but there's no need to memorize them (unless your instructor asks you to do so). Your intuition about the fundamental behavior of gases is enough to remind you about each of them, as you will see.

Let's revisit Figure 6.12, which summarizes the four variables that we can change when we are considering a gas: T, P, V, and n. If we determine these four values, then we know the condition of our gas. In this section, for each gas law, we consider an experiment in which we change one variable and see what happens to a second variable, leaving the remaining two fixed.

For our first gas law, we return to Figure 6.11 and the 22.4-liter box at STP that contains exactly one mole of gas particles. If we sit on the box and make it half the size, but keep the box sealed and at the same temperature, what happens? Of our four variables, what has changed and what has stayed the same? In this case V has changed and T has stayed the same. The variable n has also stayed the same, because we did not allow any gas to escape. But what about P? What happened to it?

When we made the box smaller, how did the environment of the gas inside it change? In the smaller box, each gas particle has a smaller distance to travel to reach a wall. That means there are more collisions with the inside surface of the box, which means the pressure of the gas must have increased. We have just deduced that when we keep T and n the same and make V smaller, we make P bigger. We can express this using an if-then statement with arrows, or we can express it mathematically as shown in **Figure 6.13**.

The symbol $\propto$ between the P and the $1/V$ on the bottom of Figure 6.13 is a *proportionality symbol*. It relates the thing to its left to the thing to its right. For example, imagine that that the volume, V, is big and then gets smaller, as illustrated

If $V\downarrow$ then $P\uparrow$

or

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

Figure 6.13 The Relationship between Volume and Pressure

This figure illustrates the relationship between two variables that dictate the behavior of gases. We can use an if-then statement and up/down arrows (top) or a mathematical relationship (below) to show a relationship between variables.

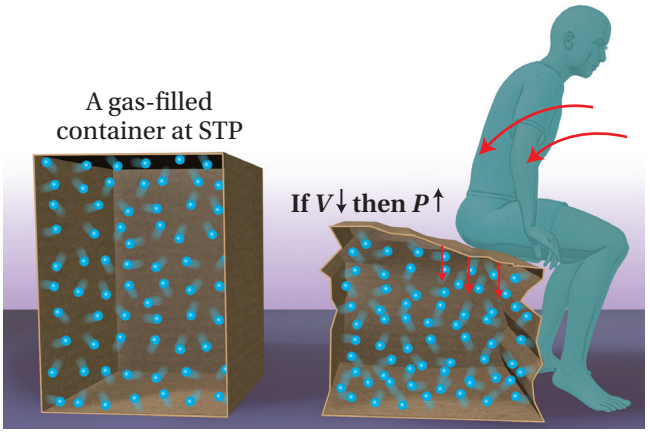


Figure 6.14 Boyle's Law

Reducing the volume of a gas-filled container, like a box, increases the pressure of the gas. Notice that the box contains equal numbers of gas particles before and after it is compressed. The compressed box contains gas at a higher pressure because the particles collide with the walls of the box more often. As volume goes down, pressure goes up. In other words, pressure is inversely proportional to volume.

in **Figure 6.14**. According to our proportionality relationship, when V gets smaller, P gets bigger. And because V and P vary in opposite directions, we say that V and P are *inversely proportional*: as one goes up, the other goes down.

This principle about the behavior of gases is known as **Boyle's law,** after Robert Boyle, a seventeenth-century British scientist. It's not necessary to memorize Boyle's law, because the relationship between volume and pressure is already second nature to us. We already know, for example, that an inflated balloon compresses when we sit on it. Compression means the balloon volume decreases, and we may worry that the balloon will burst because we sense that its pressure has increased due to the decrease in volume.

QUESTION 6.6 If you take the sealed box in Figure 6.14 and somehow make it bigger, what happens to the pressure of the gas inside the box?

ANSWER 6.6 The pressure inside the box decreases. Why? Because you have increased the volume of the box. When the volume of the box increases, the surface area also increases and there are fewer collisions per area inside the box.

For more practice with problems like this one, check out end-of-chapter Question 40.

FOLLOW-UP QUESTION 6.6 True or false? Pressure inside a container depends not only on the number of collisions in the box, but on the number of collisions per area of the inside of the box.

If we change the number of moles of a gas, the volume of the gas changes.

Our second gas law is named for Avogadro, the same man whose number defines one mole. **Avogadro's law** explains how a change in the numbers of atoms and molecules—given by the variable n—affects the volume of a gas. To understand this law, let's imagine that we have a cylinder capped with an airtight piston. The cylinder contains a small amount of gas and also has an inlet valve through which we can add gas in much the same way that air is added to a tire (**Figure 6.15**). The piston is in a position some distance up from the bottom of the cylinder because the gas molecules are colliding with the piston. These collisions keep the piston from falling to the bottom.

If we add more gas to the cylinder, more gas molecules are hitting the interior surface of the cylinder, and the piston is forced upward. As the piston ascends, the volume of the cylinder increases. When the pressure inside the cylinder equals the pressure of the atmosphere pushing down on it, the piston stops rising. Thus, in this experiment, the variables T and P are held constant and the variables V and n are changing. We can describe this relationship with these words: As the number

of moles of gas increases, so does the volume. We can write this relationship using symbols and a proportionality symbol ($\propto$):

$$V \propto n$$

where n represents the number of moles of gas, and V represents the volume of the gas. Recall that the proportionality sign tells us that as one quantity goes up, so does the other. Likewise, when one quantity goes down, the other does as well.

As we all know from personal experience, Avogadro was right in saying that the volume of a gas increases when more gas is added to a container. For example, we know that when we blow up a balloon, the balloon gets bigger (**Figure 6.16a**). We know that when we push air from our lungs into our cheeks while keeping our mouth shut, our cheeks puff out (**Figure 6.16b**). We know that when we inhale air into our lungs, our chest cavity increases in size (**Figure 6.16c**). It's second nature to us: as the number of moles of gas increases, volume also increases.

Notice this common thread for the gas laws: they all involve a change in only two of the four variables. In Figure 6.15, the number of moles of gas is increased by forcing gas molecules into the cylinder. Volume is the only other thing permitted to change, and the experiment is designed to tell us what happens to the volume (second variable) as the value of n (first variable) increases. The other two variables

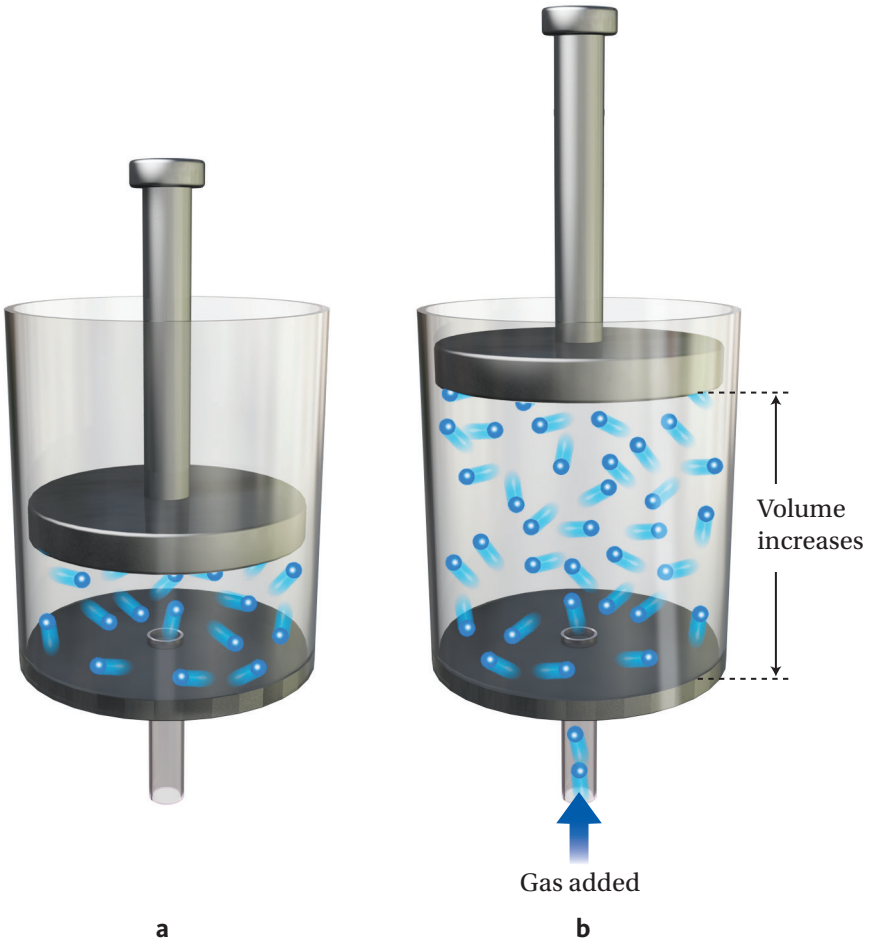


Figure 6.15 Gas Volume and Amount of Gas

(a) In a cylinder fitted with a piston, the pressure of the gas molecules determines the height of the piston and therefore the volume of the cylinder. (b) When more gas molecules are added to the cylinder, the piston rises and the volume increases.

Figure 6.16 Avogadro's Law in Action

In any flexible container, gas volume is proportional to the number of moles of gas present. (a) When you blow up a balloon, you are increasing the number of moles of gas (air) it contains. The volume of the (flexible) balloon increases. (b) Your cheek volume increases when the number of moles of gas (air) in your (flexible) mouth increases. (c) When you inhale a larger-than-usual amount of air, the volume of your (flexible) lungs increases your chest size noticeably.

depicted in Figure 6.15, pressure and temperature, are not permitted to change, and we say that they are *held constant*.

QUESTION 6.7 A bicycle tire runs over a tack and deflates. How do the four variables that we usually keep track of in gas experiments change when this occurs?

ANSWER 6.7 As the tire deflates, it releases moles of pressurized gas. Thus, n and P decrease. V decreases, too, because the tire gets smaller, but T will not change substantially as the tire deflates.

For more practice with problems like this one, check out end-of-chapter Question 39.

FOLLOW-UP QUESTION 6.7 To test Avogadro's law for yourself, you decide to use a bicycle pump to force air into a metal box that has a gas inlet valve. What is wrong with your plan? What error have you mistakenly introduced into your experiment?

If we change the temperature of a gas, the volume or pressure changes.

Let's return to our rigid, tightly sealed box from Section 6.3 that has a volume of exactly 22.4 liters. As you recall, at STP our box contained exactly one mole of gas particles. Imagine that at standard temperature, 0°C, the gas molecules inside the box are moving at some average speed. If we increase the temperature in the box, what happens? As we increase the temperature, the average speed of the gas molecules increases. This means that each molecule collides with the inside surface of the box more often than it did at 0°C. Because the box is rigid and cannot expand, the volume remains constant. What does change is the pressure, because the number of collisions on the inside surface of the box has increased. We can see that as the temperature increases, the pressure increases, too. Pressure and temperature are proportional to one another:

$$P \propto T$$

We can describe this event in words: *If the volume of a gas and the number of moles of gas are fixed, an increase in gas temperature results in an increase in gas pressure.* This relationship is sometimes called **Amontons' law** after the scientist who first studied it.

Let's change things a bit. Imagine now that the box is flexible, like a cube-shaped balloon. In this case, when the temperature increases, and the number of collisions increases, the box gets bigger. And, because we are allowing the box to get bigger, the pressure is held constant. P and n stay the same while V and T change. We can say that as the temperature increases, the volume also increases. The opposite must also be true: if we lower the temperature, the volume of the cube will decrease. We can write this as a proportion:

$$V \propto T$$

We can also state this in words: *When the temperature of a gas is increased, the volume also increases.* This relationship is sometimes called **Charles's law** after the

scientist who first studied it. These two laws, in which temperature is changing, are illustrated in **Figure 6.17**.

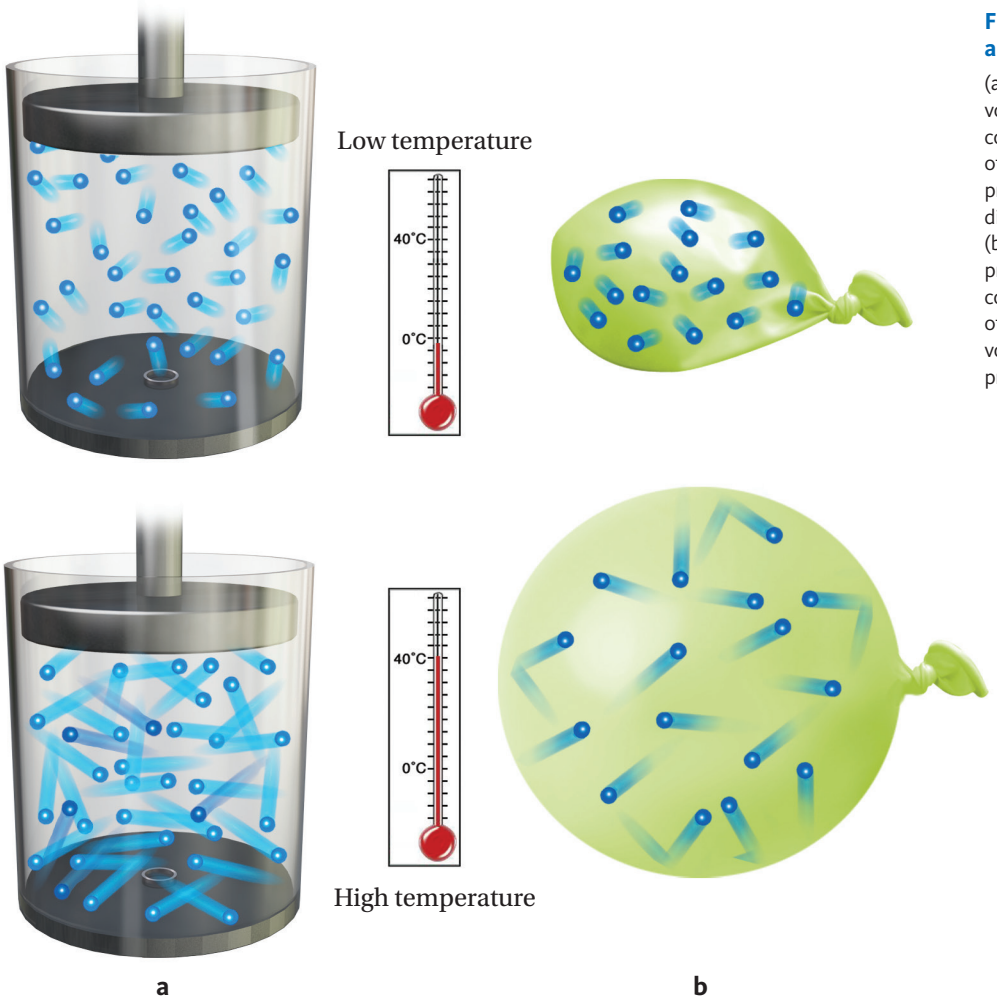


Figure 6.17 Charles's Law and Amontons' Law

(a) Amontons' law tells us that with volume *V* and number of moles *n* held constant, increasing the temperature of a gas-filled container increases the pressure of the gas. Pressure is directly proportional to temperature. (b) Charles's law tells us that with pressure *P* and number of moles *n* held constant, increasing the temperature of a gas-filled container increases the volume of the gas. Volume is directly proportional to temperature.

QUESTION 6.8 What happens to the mean free path of the molecules in a sample of gas when the volume of the gas is decreased and the temperature is held constant?

ANSWER 6.8 When the volume decreases as the temperature stays unchanged, a molecule collides with the walls of the container and with other molecules more frequently. The mean free path gets shorter.

For more practice with problems like this one, check out end-of-chapter Question 47.

FOLLOW-UP QUESTION 6.8 You decide to buy your cousin a balloon for her birthday, which is in February. Leaving the warm store, you walk out into subzero temperatures. Which one of the four gas variables does not change significantly when you walk outside?

The four gas laws are summarized in **Figure 6.18** (on p. 178).

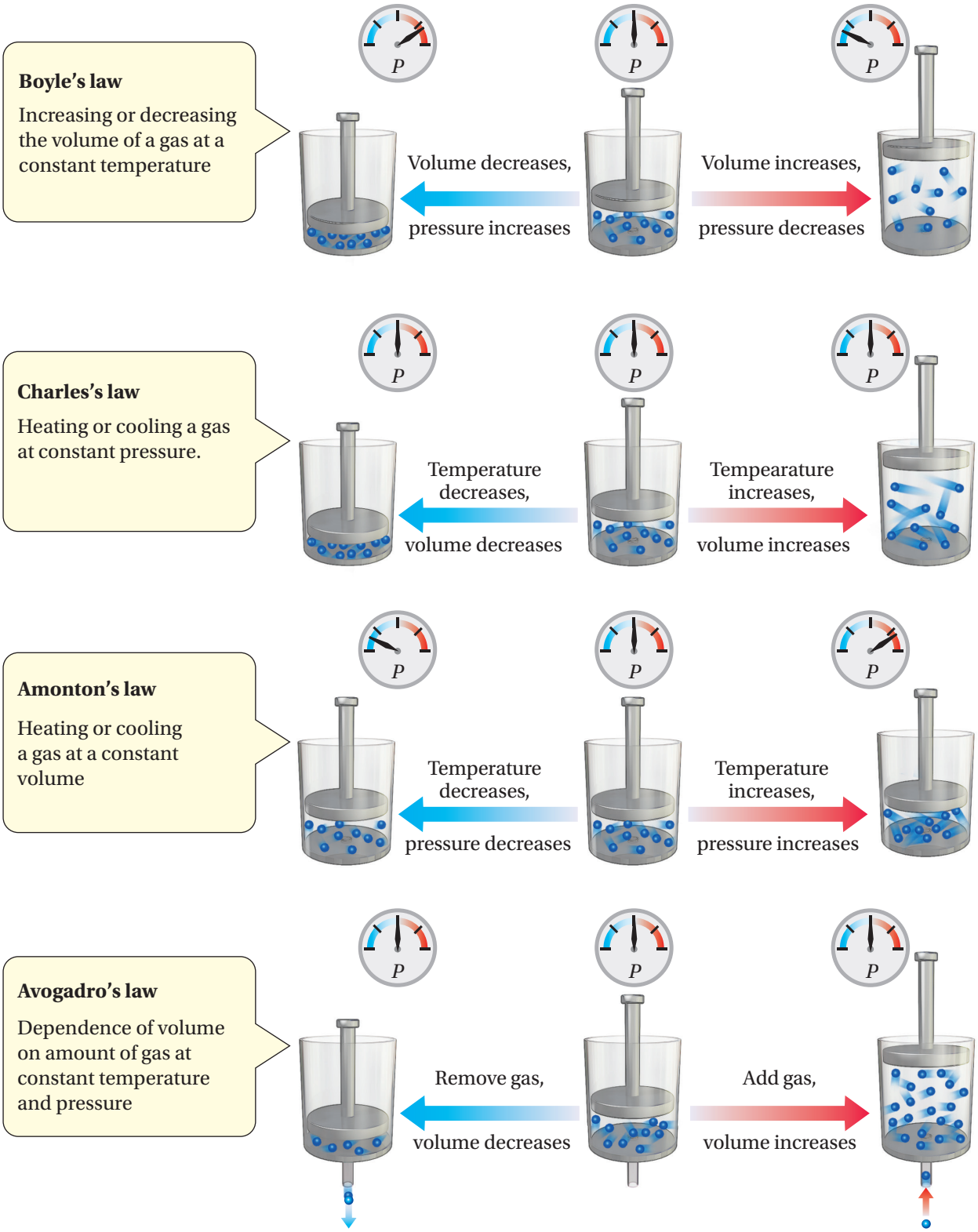


Figure 6.18 A Visual Summary of the Gas Laws

TOP STORY Bees as Gas Detectives

Bees are devoted busybodies. Because they are small—a typical bee weighs about half a gram—they can fly in and out of tiny nooks and crannies, all the while interacting with molecules in the air. They can take those molecules back to the hive along with the pollen, the dusty particles from certain plants, they collect. The exact number of particles a bee encounters on a flight depends on gas-law variables such as temperature and pressure. But because the mean free path of the particles in air is very short, on the order of nanometers, any bee encounters particles constantly as it flies.

Figure 6.19 How Bees Bring the Local Environment Back to the Hive

(a) A microscopic view focuses on barbed bee hairs, which catch particles in air and pollen from flowers. (b) This bee is deep inside a flower and coated with pollen, which it takes back to the hive along with other particulates.

The hairs on a bee are specially designed to attract pollen, hold on to it, and transport it back to the hive. Dust particles in the air naturally stick to those bee hairs as they fly, and those particles can contain molecules of any substances found in the air. In fact, bees have been described as "flying dust mops." And, because bees are breathing while they are flying, their bodies take in the surrounding air and the water contained in that air. They bring those air and water samples back to the hive in their bodies. In this way, bees serve as a means of sampling water, gases, particulates, and plant matter in the area around the hive they call home. **Figure 6.19a** shows a micrograph of jagged bumblebee hair, and **Figure 6.19b** shows a bee happily headfirst in a yellow flower.

When bees return to the hive, they beat their wings furiously to cool off the temperature inside the hive. This beating motion releases the molecules that were stuck to the bees' bodies and hair. The molecules then become concentrated in the hive space. If a bee takes nectar or pollen from a plant containing a foreign molecule of some sort, or if the bee flies through air contaminated with some foreign molecule, those molecules eventually will end up in the hive, sometimes in the honey and at other times in the bees themselves.

Scientists like bee expert Jerry Bromenshenk of the University of Montana are trying to exploit the natural talent that bees have for taking home samples of the air they fly through. Bromenshenk and his research group have taught bees to seek out specific non-nectar molecules, and he says this task is easier than training a bloodhound to follow a scent.

How can this work possibly be relevant to humans? One application is in the detection of land mines, which sometimes contain the explosive molecule trinitrotoluene (TNT). After being trained to sniff out TNT molecules (**Figure 6.20**), bees are amazingly adept at locating buried land mines. Bees can be trained to fly to spots where they smell the explosive compound TNT. The bees swarm in the areas highest in TNT molecules, which are the places where the land mines are located. In countries where land mines are prevalent, such as Cambodia, Somalia, and Angola, mine removal is an urgent priority. With luck, the coming years will see mine-tracking bees working in these locations.

So why use bees rather than dogs, which are known to be great mine-sniffers? The answer is that although both bees and dogs have great senses of smell, bees are easier to train. In fact, they can be taught to sniff out an explosive in about two days. And, because bees fly solo and don't need to be on a leash with a handler, concerns about a human or a dog tripping a land mine are allayed when bees are used instead. Bees are also being trained to locate dead bodies and chemical or biological warfare agents, other jobs that dogs have traditionally performed.

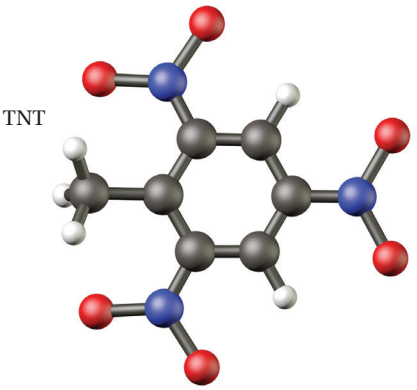


Figure 6.20 The Hunt for TNT

TNT is an organic molecule with the chemical formula $C_7H_5N_3O_6$. (Black balls = carbon atoms; red = oxygen atoms; white = hydrogen atoms; blue =nitrogen atoms.)

Chapter 6 Themes and Questions

recurring themes in chemistry chapter 6

recurring theme 2 Things we can see with our eyes contain trillions upon trillions of atoms. Atoms are much too small to view with the naked eye. (p. 171)

A list of the eight recurring themes in this book can be found on p. xv.

ANSWERS TO FOLLOW-UP QUESTIONS

Follow-Up Question 6.1: (a) liquid (b) gas (c) solid. *Follow-Up Question 6.2:* False. They collide with both. *Follow-Up Question 6.3:* They are less dense on Mount Whitney and denser at the beach. *Follow-Up Question 6.4:* (b) joules. *Follow-Up Question 6.5:* One hundred times Avogadro's number, which is the same as adding two to the exponent: 6.02×10^{25} gas particles. *Follow-Up Question 6.6:* This is true. *Follow-Up Question 6.7:* To test Avogadro's law, the box must be flexible so that volume (V) can change. *Follow-Up Question 6.8:* The variables P, V, and T will decrease. The variable n will stay the same.

End-of-Chapter Questions

Questions are grouped according to topic and are color-coded according to difficulty level: **The Basic Questions***, **More Challenging Questions****, and **The Toughest Questions*****.

The Nature of Gases

***1.** An odor brought into a room by a very smelly person takes a while to travel from that person to your nose. Imagine this scenario on a hot day and a cold day. Does the travel time differ on those two days? On which day does the scent travel faster? Why?

***2.** Imagine you have a 2-L sealed flask containing 1 mole of xenon atoms. You measure the average speed of these atoms to be 700 km/h. Is every atom of xenon in the sample moving at this speed? Explain.

***3.** Is the following statement true or false? All gases mix together completely regardless of the types of atoms or molecules they contain. This is true because the atoms or molecules in a gas are all very far apart from one another and interact only rarely during collisions.

****4.** Your roommate is becoming more and more malodorous every day. You decide to adjust your thermostat to try to slow down the movement of her "fragrance" through your apartment. Will you turn on the heat or the air conditioning? Explain your answer.

**5. Using your knowledge of gas behavior, explain why, during a chemical weapon attack, people farther from the gas source have a greater chance of survival.

**6. True or false? Gases are not susceptible to the force of gravity. Explain your choice.

**7. Referring to Figure 6.4, explain what is different among the three balloons shown. Which has the greatest pressure? Which has the greatest volume?

**8. Is the following statement true or false? If air sample A is denser than air sample B, then the mean free path of the particles in gas air sample A is shorter than the mean free path of the particles in air sample B.

***9. For each substance described here, indicate whether it is a solid, liquid, or gas at room temperature. (a) a green substance with a melting point of 234°C (b) a clear substance that adopts the shape of the container you put it in but does not fill it completely (c) a substance that maintains its shape when you remove it from the box that held it (d) a substance that disappears immediately when you uncork the bottle that contains it

***10. Figure 6B shows the pheromone molecule of the light brown apple moth (LBAM). Write the chemical formula of this molecule.

Air and Atmospheric Pressure

*11. Which gases are not found in significant quantities in air?

(a) mercury gas (b) nitrogen gas (c) carbon dioxide gas (d) krypton gas

*12. In your own words, describe what is meant by atmospheric pressure. Name a pressure unit used in meteorological work.

*13. Use your understanding of the material in Chapter 6 to answer this question: Which types of matter might you find in air?

(a) a pheromone molecule (b) a dust particle (c) a diamond

(d) a nitrogen molecule (e) a piece of granite

**14. To keep your shampoo bottle the same size during your trip from San Francisco to the Rocky Mountains, you decide to remove air from the bottle as you pack it in San Francisco so that it collapses inward. In the mountains, you blow into the bottle so that it puffs outward and cap it quickly. Is this a reasonable plan? Why or why not?

**15. Death Valley, California, is 86 meters below sea level. If you pack a plastic tube of sunscreen at the beach in San Diego and drive to Death Valley, will the tube be larger or smaller than it was in San Diego?

**16. How do bees concentrate air pollutants in their hive? What practical uses have been found for bees as a result of this ability?

**17. In your own words, explain why bees are preferred over dogs as mine-sniffers.

***18. According to Table 6.1, how many molecules of nitrogen are there in a sample of air that contains the following?

(a) a combined total of 100,000 atoms and molecules

(b) 1 mole of gas

***19. The barometric pressure is 730 mm Hg on a low-pressure day in Hooverville. Express this pressure in the following units:

(a) inches of mercury (b) atmospheres

Pressure and Units Used to Report It

*20. In your own words, describe the difference between force and pressure. What does the word *pressure* mean when it refers to a gas?

*21. In your own words, explain why it is possible to express atmospheric pressure with a distance unit rather than a pressure unit.

*22. In your own words, explain why millimeters of mercury is the unit often used for barometric measurements.

**23. Sketch a barometer and label its parts. In your own words, explain how a barometer is used to measure atmospheric pressure.

**24. If the pressure at sea level is approximately one atmosphere or 760 mm Hg, which of these pressure measurements could have been made on the summit of Mount Everest?

(a) 4.5 atm (b) 1000 mm Hg (c) 0.30 atm (d) 1.2 atm

**25. The pressure unit of pounds per square inch illustrates the nature of pressure because it shows a mass pushing on a square surface. Which of the following imaginary units would not be a unit of pressure?

(a) pounds per square centimeter (b) pounds per meter

(c) meters per square second

***26. Consider the two clear boxes shown in Figure 6.9. Does one contain more oxygen molecules than the other? Does one have a higher percentage of oxygen molecules than the other?

***27. Mercury, a very dense liquid, is the most common liquid used to make barometers. Speculate on why water would not be a good choice to replace mercury. How would a water-based barometer look different from a regular, mercury-based barometer?

Moles and Molar Volume

*28. How many gas particles are contained in 100 moles of gas?

*29. What are the values of the four gas variables for a container at STP that has a volume equal to the molar volume?

*30. What does the abbreviation STP mean?

**31. Which of these objects would likely *not* be counted using moles?

(a) one-carat diamonds (b) atoms of argon

(c) molecules of carbon dioxide (d) sodium ions

**32. If a flexible container at STP holds 1.2×10^{23} molecules of nitrogen, what is its volume?

**33. Suppose you have a box with a volume of 22.4 L that is at STP. This box is somewhere at sea level, where the pressure is 1.0 atm, standard pressure. Now you have another box that is the same size but is at an altitude of 10,000 feet. The second box contains a sample of air from that high-altitude environment which is at standard temperature. In the two boxes, which of the four gas variables are the same, and which are different? For those that are different, how do they change between sea level and high altitude?

**34. For each pair of gas variables, describe an experiment that would allow only the following two variables to change:

(a) P and V (b) n and P

**35. For each pair of gas variables, describe an experiment that would allow only the following two variables to change:

(a) n and V (b) V and T

***36. What is the approximate volume of a balloon at STP that contains 9.00×10^{23} atoms of helium gas?

The Gas Laws and Gas-Law Variables

*37. In the imaginary lakeside town of Goolikobruquik, the price of potatoes always goes up when the temperature of the lake water goes down. When the water temperature goes up, the price of cherries inevitably goes up. For each pair of variables, indicate whether they are directly proportional or inversely proportional to each other.

*38. In your own words, explain what is meant by the terms *inversely proportional* and *directly proportional.* Give one example of each relationship from this chapter.

*39. Which of these pairs of gas variables are inversely proportional to one another?

(a) P and V (b) n and V (c) V and T (d) T and P

*40. A rigid container has a piston that can move up and down freely. If the temperature of the gas in the cylinder is increased, what other gas variable will change?

*41. A box is sealed and rigid. If you add gas to the box through a valve, which variables are changing? Which variables do not change?

*42. On a hot summer day, a car tire bursts when the temperature inside it increases. Which of the gas laws does this scenario describe?

**43. A kid pumps up his bike tire, and it becomes very firm. If the temperature is assumed to stay constant, which of the four gas variables change when the tire is pumped up? Which do not change?

**44. A balloon is heated by a candle flame held at a distance to prevent the balloon from melting. It pops. If you assume that the balloon stays the same size until just before it pops, which variables change when the balloon is heated? Which do not change?

**45. A window breaks in an airplane traveling at 32,000 feet. Does the pressure within the plane increase or decrease when this happens? Explain.

**46. A container fitted with a movable piston is used to demonstrate Avogadro's law. As gas is introduced into the cylinder, the piston moves upward and the volume increases. Why does the pressure of the gas remain constant during this experiment?

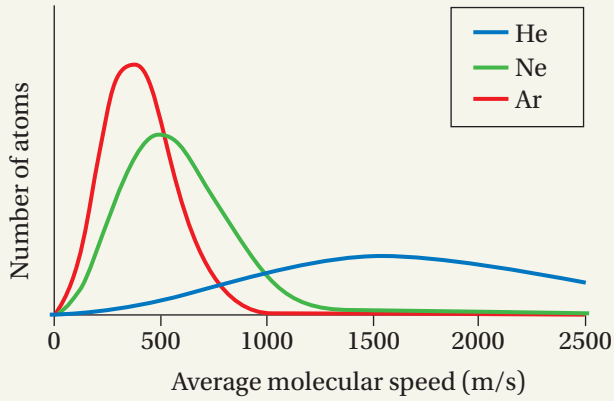
**47. For each of these situations, indicate which gas particles have the longer mean free path.

(a) 2 billion gas particles in a 1-L container or 4 billion of gas particles in the same container

(b) 4 billion gas particles in a 4-L container or 4 billion gas particles in a 2-L container

**48. A rigid, sealed box contains two molar volumes of gas at STP. What is the size of the box? If you want to increase the pressure in the box, which gas variable could you change?

***49. A box at 0 degrees Celsius has a volume of 11.2 liters and contains one-half mole of gas particles. The gas particles are a 50–50 mixture of nitrogen gas (N_2) and oxygen gas (O_2). What is the pressure in the box, expressed in millimeters of mercury?



***50. This graph shows the average molecular speeds of three gases at 27°C.

(a) Use the periodic table to rank these gases in order of increasing atomic mass.

(b) What correlation do you observe between molecular speed and molar mass?

(c) Estimate the average speed of a neon atom at 27°C.

*51. In this chapter, we discuss the idea of gas pressure and see what can happen to a shampoo bottle that is taken from a place with high atmospheric pressure to a place with low atmospheric pressure. But while the place where we live is at a specific atmospheric pressure (although that may vary a bit with changing weather), there are things in your environment that may be at higher or lower pressure. Imagine the following objects and predict if the pressure in that object will be greater than or less than the atmospheric pressure.

(a) a new can of Fix-a-Flat on a garage shelf

(b) a bottle of vitamins bottled and shipped from a location at very high altitude

(c) a piece of fish vacuum-sealed in a plastic bag

(d) a can of whipped cream

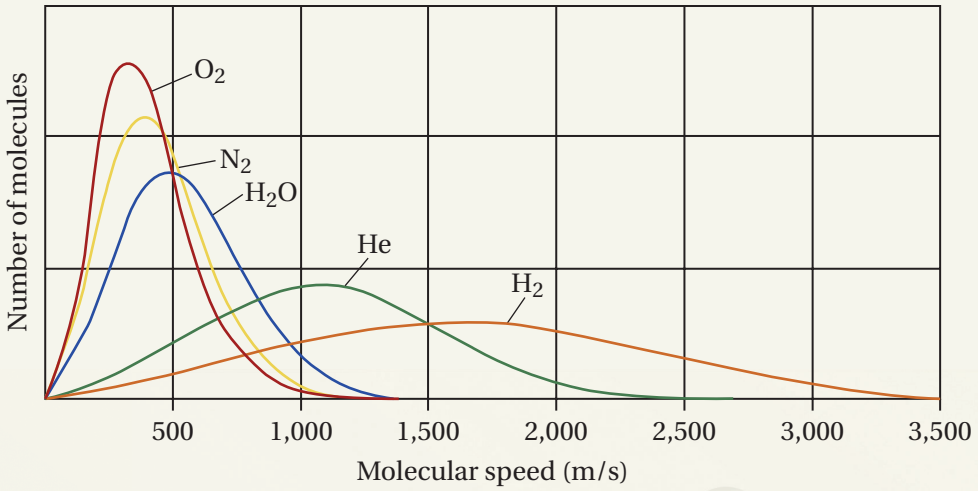
*52. The tire pressure on your bike or your car's tires will vary with changes in temperature throughout the year. Would you expect your tire pressure to be higher or lower in winter versus summer? Tire pressure can also change as the tire is used. Would you expect tire pressure to increase or decrease as you travel on the tire?

**53. If you have a gas stove that uses natural gas, then you probably have a pipe leading into your home that delivers it. These pipes, which typically are 20 to 42 inches in diameter, are built to withstand pressures greater than what is needed. The pressure of natural gas in municipal pipelines is typically in the range of 200 psi to 1,500 psi. "Psi" is a nonmetric unit for pressure that stands for "pounds per square inch." It can be converted to pascal (P), the metric unit for pressure, using this conversion:

$$1 \text{ psi} = 6{,}894 \text{ P}$$

If the pressure of natural gas in a municipal pipeline is four million pascals, what is the pressure expressed in psi?

***54. Figure 6.2 in Section 6.1 shows the average speed of a collection of nitrogen molecules at a given temperature. That graph shows that at 27°C, the typical nitrogen molecule moves at a speed of about 500 meters per second (m/s). The graph below shows the same speed of nitrogen molecules at the same temperature but also shows several other atoms and molecules as well. Examine this graph and try to answer this question: Why are some atoms or molecules moving faster than others? Here is a hint: consider the masses of each atom or molecule. Is there a relationship between the mass and the speed? Does this trend make sense to you?



55. INTERNET RESEARCH QUESTION

Return to the Nature Box in this chapter entitled, "Is Natural Gas the Ideal Energy Resource?" In this essay, the dramatic rise of hydraulic fracturing in Pennsylvania, Ohio, and West Virginia is discussed. Earthquakes due to hydraulic fracturing in Oklahoma, Arkansas, and Colorado are also mentioned. These states all have underground deposits of natural gas that are being tapped by hydraulic fracturing, but there are many other states where hydraulic fracturing is used, or could be used, to produce natural gas.

Is hydraulic fracturing being used in your area? To find out, go to the website of the U.S. Energy Information Administration. Under the Sources and Uses tab, select Natural Gas. Now select the U.S. Natural Gas Infrastructure Map. Under "Find address" enter your zip code or full street address.

Answer these questions about natural gas drilling and use in your area.

(a) Are there power plants in your area that run on natural gas?

(b) Do you live over an area with shale gas, colored in tan on the map?

(c) Do you see blue dots in your area on the map? These dots represent gas well locations.

(d) If you don't see any of these features, zoom out away from your local area until you do. What do you see? How far away are the nearest gas power plants, the nearest shale deposits, or the nearest gas wells?