CHAPTER TEN

Solids, Liquids, and Gases

Figure 10.1 (a) Coal is the solid form of fossil fuels. It must be mined. (*continued*)

The North Dakota Boom

In 2006, geologists discovered an enormous underground reserve of oil and natural gas in western North Dakota. The impacts of the discovery were huge: While much of the global economy staggered, the Dakotas boomed, creating thousands of jobs and enormous wealth. In 2006, North Dakota was ranked 40th among U.S. states for per capita income. By 2013, it was ranked second.

Coal, oil, and natural gas are collectively called *fossil fuels*. Composed mainly of carbon and hydrogen, these fuels form by the decay of plant and animal matter over long periods of time. Along with nuclear energy, fossil fuels are the workhorse of the industrialized world: Each year, fossil fuels produce over 80% of the energy used in the United States.

Among fossil fuels, natural gas has steadily grown in importance. The surplus of oil and gas coming out of North Dakota and Montana has pushed down the cost of fuel, which in turn has reduced the cost of other commodities. Your gas bill and your grocery bill are lower because of the massive 2006 discovery. And though all fossil fuels produce carbon dioxide, natural gas burns more cleanly and efficiently than coal or oil, reducing the environmental impact of energy consumption. Natural gas is widely used for heating, and many vehicle fleets (like buses, cabs, and delivery trucks) have transitioned from liquid gasoline to natural gas.

Collecting and transporting each type of fossil fuel has its own challenges. Coal, a solid, has to be mined and then transported by ship or train. Companies extract oil (a liquid) and natural gas from wells and often transport these commodities through massive pipelines (**Figure 10.1**).

Figure 10.1 (continued) (b) Oil is the liquid form of fossil fuels. (c) An operating oil well. (d) A worker manages natural gas lines. (e) Many newer buses run on natural gas. This fuel burns more cleanly than gasoline. (a) Vyacheslav Svetlichnyy/Shutterstock; (b) Lowell Georgia/Getty Images; (c) Ed Reschke/Stone/Getty Images; (d) ZoranOrcik/Shutterstock; (e) George Rose/Getty Images

Safely handling and transporting gases can be especially challenging. Workers must follow special safety guidelines when working with natural gas. In fact, many careers, from health care to mechanical work to restaurant management, involve work with compressed gases. Understanding how gases behave will serve you well in nearly any field.

In this chapter, we'll examine the properties of matter in its three phases: solid, liquid, and gas. We begin with solids and liquids, exploring how chemical bonds affect a substance's structure and properties. Then we'll shift our attention to gases. We will see how the pressure and volume of gases vary with temperature and with the amount of gas present. We'll explore chemical processes that involve gases, such as combustion reactions or the production of beer.

At the end of this chapter, we will return to the challenge of safely storing and transporting natural gas. We'll look at new discoveries that are changing the way scientists and engineers think about this problem and opening up exciting opportunities for the future.

Intended Learning Outcomes

After completing this chapter and working the practice problems, you should be able to:

10.1 Interactions between Particles

- Describe the motion of particles in a solid, liquid, or gas.

10.2 Solids and Liquids

- Describe the bonding and arrangement of particles in ionic, metallic, molecular, polymeric, and covalent-network substances.
- Describe the different types of intermolecular forces, and relate these differences to relative melting or boiling temperatures.

10.3 Describing Gases

- Describe the key features of an ideal gas.
- Describe how to use a liquid barometer to determine pressure.

10.4 The Gas Laws

- Apply the combined gas laws to relate changes in the pressure, volume, and temperature of a gas.
- Relate the pressure, volume, number of moles, and temperature of a gas using the ideal gas law.
- Relate the temperature, volume, and pressure of a gas to atomic or molecular motion.

10.5 Diffusion and Effusion

- Describe the motion of larger and smaller gas particles at a given temperature, and apply these concepts to the principles of diffusion and effusion.

10.6 Gas Stoichiometry

- Apply the principles of stoichiometry to solve problems involving reactions of gases.

10.1 Interactions between Particles

In Chapter 1, we discussed the three common *states of matter* (also called the *phases of matter*): solid, liquid, and gas. We saw that the macroscopic (human-scale) properties are a function of their structure on the atomic or molecular level. In a solid, particles pack closely together. Each atom is held in a fixed position by the atoms around it. In a liquid, the particles are close together, but they are not held in a fixed position; particles move freely past each other. In a gas, the particles move independently of each other—they are spaced far apart and have little or no interaction with the other gas particles as they move. These concepts are summarized in **Table 10.1**.

TABLE 10.1 The States of Matter

	Atomic/Molecular Arrangement	Macroscopic Properties
Solid	Particles are close together and held in a fixed place.	Definite shape and volume
Liquid	Particles are close together, but move freely past each other.	Definite volume Adopts shape of container
Gas	Particles are far apart and have very little interaction.	Adopts shape and volume of container

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The melting and boiling points of a substance depend on the forces holding the particles together. The stronger the forces of attraction, the higher the phase transition temperatures.

Substances with strong forces of attraction have high melting and boiling points. ■

A transition from one state of matter to another is called a *phase change*. For example, let's look at the phase changes that take place with water: In ice, the water molecules pack tightly together (**Figure 10.2**). If we add heat (kinetic energy) to the ice, the molecules in the solid vibrate more rapidly. If we continue to add heat, the molecules gain enough energy to overcome the forces holding them in place. They break out of the solid framework and begin to move freely past each other. The ice melts—a phase change from solid to liquid. If we then heat the liquid water, the molecules move faster and faster until they begin to break out of the liquid phase and enter the gas phase. Particles in the gas phase move about freely, interacting very little with each other.

For a compound to change from solid to liquid or from liquid to gas, there must be enough kinetic energy to overcome the forces of attraction. The stronger the forces between the particles, the greater the energy required to change

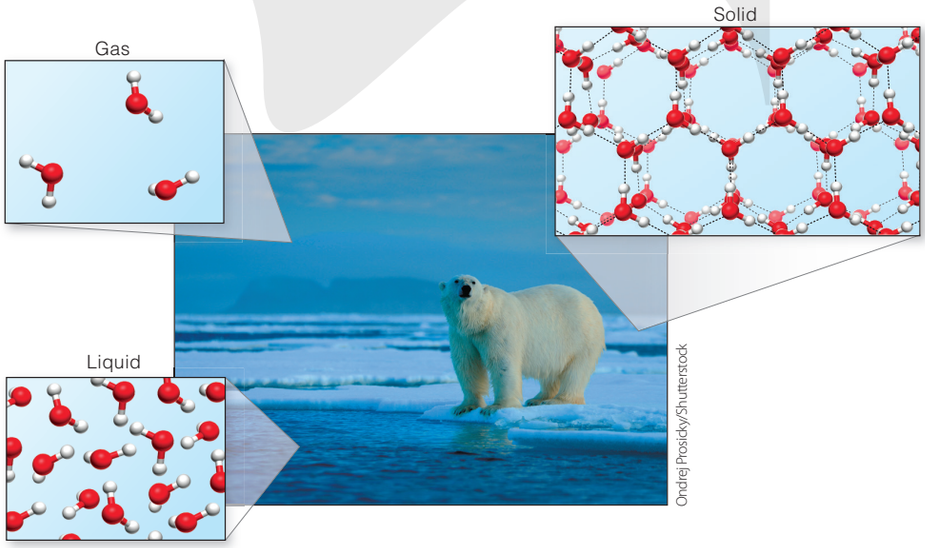

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Figure 10.2 This image captures water in all three phases: solid ice, liquid water, and gaseous steam. Each phase change requires heat energy to overcome the forces of attraction between the water molecules.

Explore
Figure 10.2

phase. A substance with very strong forces between its particles will have a high melting point and boiling point. We can use phase transition temperatures to understand and compare the forces that hold substances together.

TRY IT

1. Methane, ammonia, and water are all covalent compounds that form molecules. Based on the melting points shown, which compound exhibits the strongest forces of attraction between molecules? Which one exhibits the weakest forces of attraction?

Substance	Formula	Melting Point
Methane	CH_4	–182 °C
Ammonia	NH_3	–78 °C
Water	H_2O	0 °C

Check it
Watch Explanation

10.2 Solids and Liquids

In solids and liquids, forces of attraction hold particles close together. These forces vary for different types of substances. In this section, we will explore the different ways that particles can be arranged within solids and liquids and learn about the properties that result from their unique structures.

Ionic Substances

We first talked about ionic compounds in Chapter 5. In these compounds, positive and negative ions pack tightly together in rigid frameworks, called *lattices* (**Figure 10.3**). Because the interactions between the ions are so strong, it takes a tremendous amount of energy to disrupt the lattices. As a result, most ionic compounds have very high melting points (**Table 10.2**).

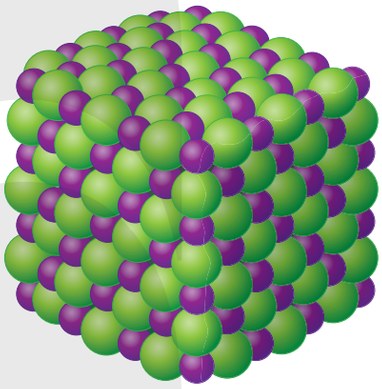

Figure 10.3 Ionic compounds form ordered lattices of alternating positive and negative charges.

TABLE 10.2 Melting Points for Ionic Compounds

Compound	Melting Point (°C)
NaCl	801
KCl	770
$MgCl_2$	714
CaO	2,572
Al_2O_3	2,072

Metallic Substances

Elemental metals (as well as mixtures of metals, called *alloys*) usually form ordered lattices of tightly packed atoms (**Figure 10.4**). Metallic atoms have a loose hold on their outer electrons, and the electrons move easily from one atom to the next. Because of this property, metals are good conductors of electricity.

While atoms pack close together in metallic solids, the neutral metal atoms do not bind as tightly to each other as the oppositely charged ions in an ionic compound. Because of this, a metal lattice is not as rigid as an ionic lattice, and the shapes of metallic solids are easily altered. Metals are *malleable*, meaning they can be pounded into different shapes, and they are *ductile*, meaning they can be stretched into wire.

Metals have moderate to high melting points (**Table 10.3**). Metals like iron, aluminum, and copper are commonly melted and molded into useful shapes (**Figure 10.5**).

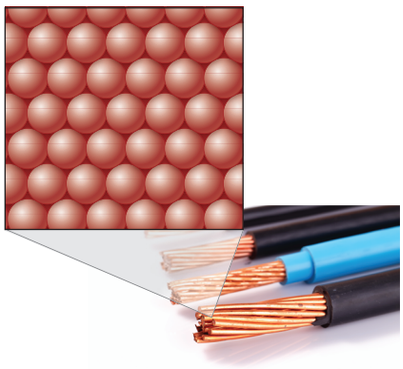

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Figure 10.4 Metals pack into close-fitting arrangements of atoms.

TABLE 10.3 Melting Points for Common Metals

Element	Melting Point (°C)
Lead	327
Aluminum	660
Gold	1,064
Copper	1,085
Iron	1,538

Cultura Creative (RF)/Alamy

Figure 10.5 Workers pour molten (liquid) iron into a mold.

Molecular Substances

Recall that covalent bonds form between nonmetals.

Recall that molecules are discrete units of atoms held together by covalent bonds. Covalent bonds within a molecule are very strong, but the forces *between* the individual molecules—called **intermolecular forces**—are much weaker (**Figure 10.6**). As a result, molecular compounds usually have lower melting and boiling points than ionic compounds.

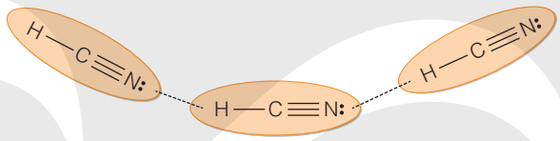


Figure 10.6 The forces between molecules (shown as dashed lines) are weaker than the covalent bonds within a molecule.

The strength of intermolecular forces varies widely, depending on the structure of the molecules. As a result, molecular materials exhibit a stunning diversity of physical properties. Broadly, intermolecular forces are divided into three key groups: *dipole–dipole interactions*, *hydrogen bonds*, and *dispersion forces*.

Intra–* versus *Inter–

Be careful when using the prefixes *intra–* and *inter–*. *Intra–* means "within." *Inter–* means "between." So "*intra*molecular forces" are the bonds within a molecule—strong covalent bonds. "*Inter*molecular forces" are the weaker forces between different molecules.

Dipole–Dipole Interactions

As we saw in Chapter 9, many molecules contain polar covalent bonds. Depending on their shape, molecules with polar bonds often have a *net dipole*, meaning different sides of the molecule have a slight positive and negative charge (**Figure 10.7**). If a compound has a net dipole, the molecules of that compound tend to "stick" together due to the attraction of the positive and negative poles (**Figure 10.8**). This type of interaction is called a **dipole–dipole interaction**.

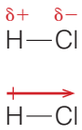


Figure 10.7 Recall that we show polar covalent bonds and molecular dipoles with the δ+ and δ− symbols, or with an arrow symbol that has a cross on the positive side and the arrowhead on the negative side.

Figure 10.8 Molecules with net dipoles experience an attractive force between their positive and negative ends.

In a polar covalent bond, the more electronegative atom has a slight negative charge.

Dipole–dipole attractions affect a substance's melting and boiling points. To illustrate this idea, consider the molecules in **Figure 10.9**. The molecules in Flask A have a molecular dipole, while those in Flask B do not. Because of the forces of

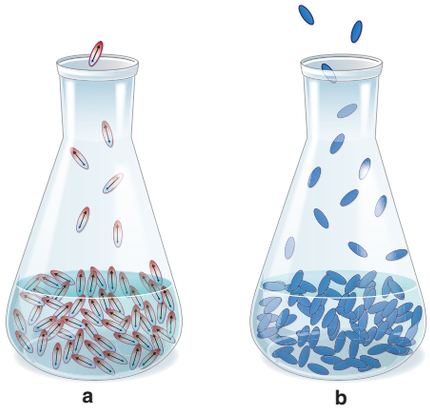


Figure 10.9 The dipoles cause the molecules in flask A to stick together more tightly, so more energy is needed for them to escape into the gas phase. As a result, molecules with a dipole tend to have higher boiling points than those that do not.

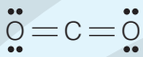
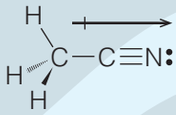
Explore
Figure 10.9

attraction between the dipolar molecules, it takes more heat energy to pull them apart. Consequently, the dipolar molecules have a higher melting point and boiling point than the nonpolar molecules have.

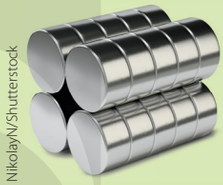
An example of this effect is shown in **Table 10.4**. Carbon dioxide has no net dipole, sulfur dioxide has a small net dipole (owing to its bent shape), and acetonitrile has a large net dipole. Notice that the larger the size of the dipole, the higher the boiling point.

Polar molecules tend to have higher melting and boiling points than nonpolar molecules. ■

TABLE 10.4 Molecules with Larger Dipoles Exhibit Higher Boiling Points

	 Carbon dioxide	 Sulfur dioxide	 Acetonitrile
Geometry	Linear	Bent	Linear
Dipole	Zero	Small	Large
Boiling Point	–79 °C*	–10 °C	+82 °C

*Carbon dioxide *sublimes* (transitions from solid directly to gas) at this temperature.


Nikolay N/Shutterstock

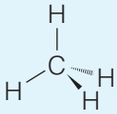
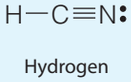
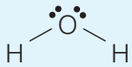
Dipole–dipole interactions are similar to the interactions between magnets—the opposite ends are attracted to each other, and they tend to stick together.

Melting and boiling points are a handy way to compare the strength of different intermolecular forces. However, we must be careful in our comparisons: Heavier molecules tend to have higher melting and boiling temperatures than lighter molecules, so it is best to compare compounds having a similar formula mass.

Hydrogen Bonding

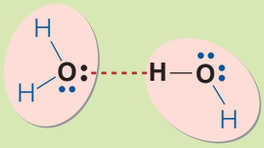
Look at the three small compounds in **Table 10.5**. The first compound, methane (CH_4), has no dipole. Because of this, CH_4 molecules have very weak interactions

TABLE 10.5 A Comparison of the Properties of Methane, Hydrogen Cyanide, and Water

	 Methane	 Hydrogen cyanide	 Water
Formula mass	16.0 u	27.0 u	18.0 u
Dipole strength*	0	2.98	1.85
Boiling point	–162 °C	+26 °C	+100 °C

*These numbers convey the relative size of each dipole.

Hydrogen bonds are stronger than other dipole–dipole interactions. ■



To form a hydrogen bond, the hydrogen atom must be covalently bonded to one electronegative atom (F, O, or N), but also attracted to the electrons on an F, O, or N atom in a neighboring molecule.

with each other, resulting in a very low boiling point (−162 °C). The second compound, hydrogen cyanide, has a large dipole: As expected, its boiling point (+26 °C) is significantly higher. But the third compound, water, is surprising: Its dipole is smaller than that of hydrogen cyanide, but its boiling point is much higher.

This surprising observation results from an especially strong type of intermolecular force called **hydrogen bonding**. Hydrogen bonding is a special dipole–dipole interaction that occurs only between molecules containing H–F, H–O, or H–N bonds. Water exhibits very strong hydrogen-bonding effects.

What causes the hydrogen-bonding effect? Recall that a hydrogen atom contains only one electron. When hydrogen bonds to a more electronegative element (F, O, or N), this one electron is pulled away, leaving the positive nucleus exposed (**Figure 10.10**). This "naked" positive charge attracts the negative poles of other molecules. Neighboring molecules with small, slightly negative atoms (F, O, or N) can get exceptionally close to the naked positive of the hydrogen atom—nearly as close as a covalent bond. The result is an unusually strong intermolecular force.

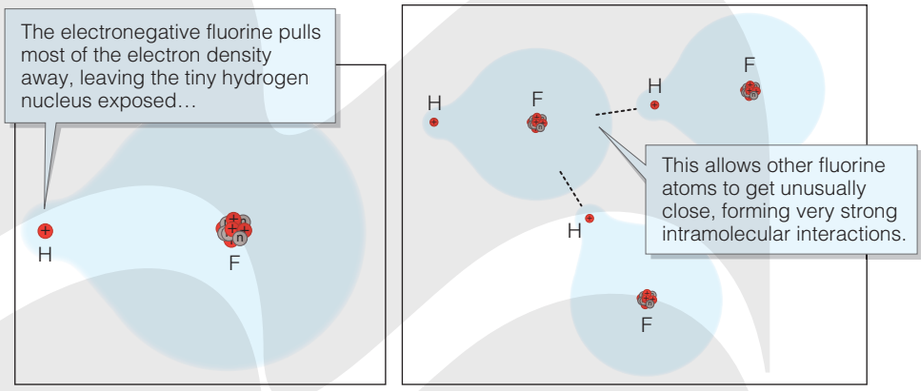


Figure 10.10 Hydrogen bonds arise when hydrogen is bonded to the electronegative elements nitrogen, oxygen, or fluorine.

Water contains two O–H bonds, resulting in multiple hydrogen-bonding interactions that cause water molecules to stick tightly to each other (**Figure 10.11**). As a result, the melting and boiling points of water are much higher than those of other small molecules. The interactions between water molecules are strong enough that some insects are able to walk on the surface of water, supported by the forces of attraction between individual molecules. The tendency of water to form droplets in the air or beads on a waxed surface also arises from this strong attraction of water molecules for each other (**Figure 10.12**). Hydrogen bonding plays a critical role in the chemistry of life, from the fundamental way water molecules interact to the shapes of huge biological molecules like proteins and DNA.

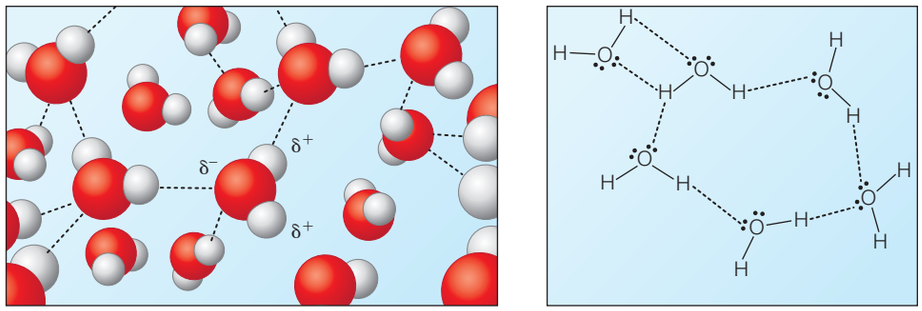
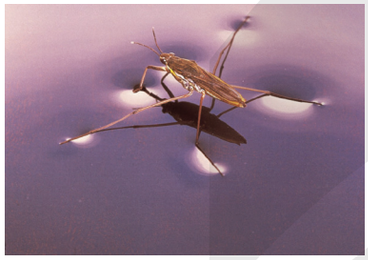


Figure 10.11 Water molecules are held tightly together by hydrogen bonds, as depicted by a space-filling model (left) and Lewis structures (right).



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a

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b

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c

Figure 10.12 Hydrogen bonding causes water molecules to stick tightly together, a phenomenon called surface tension. These forces can (a) support the weight of a bug on the water, and (b) cause water to form droplets in the air or (c) on a waxed surface. London dispersion forces are the weakest intermolecular force.

London Dispersion Forces

So what about covalent compounds that don't have dipoles? At first glance, we might expect there to be no attraction between them. However, this is not the case. Even molecules with no overall dipole exhibit a weak attraction for each other (**Figure 10.13**). These forces of attraction are called **London dispersion forces** (sometimes simply called *dispersion forces* or *London forces*). London dispersion forces are related to dipole–dipole forces but are much smaller in magnitude.

Even if they are shared evenly between two atoms, electrons are constantly moving. This motion produces slight, fleeting areas of positive and negative charges called *instantaneous dipoles*. These rapidly changing dipoles produce ripple effects in the surrounding molecules as well: A slight buildup of negative charge on one molecule will push the electron density away on a neighboring molecule. The result is a dipole on the secondary molecule and an instant of attraction between the two molecules (**Figure 10.14**).

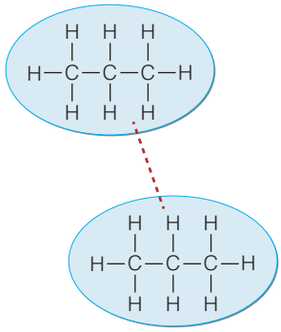


Figure 10.13 Although propane (C_3H_8) molecules do not have a net dipole, they weakly attract each other through London dispersion forces.

London dispersion forces are the weakest intermolecular force. ■

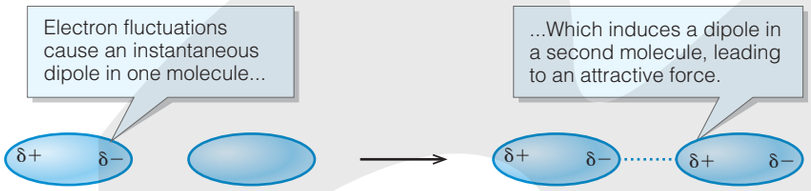


Figure 10.14 London dispersion forces are caused by the interaction of tiny instantaneous dipoles.

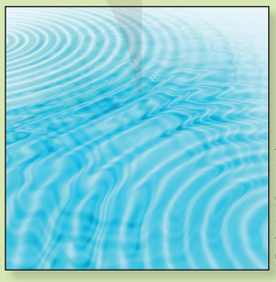
Explore Figure 10.14

Summarizing Intermolecular Forces

In summary, there are three types of intermolecular forces: London dispersion forces, dipole–dipole forces, and hydrogen bonds (**Table 10.6**). Of these, dispersion forces are the weakest—resulting in molecules with the lowest melting and boiling points. Hydrogen bonds are the strongest—resulting in molecules with the highest melting and boiling points.

TABLE 10.6 Types of Intermolecular Forces

Type	Description	Strength
Hydrogen bonding	Molecules with H–F, H–O, or H–N bonds	Strongest
Dipole–dipole forces	Molecules with net dipoles	
London dispersion forces	All covalent molecules	Weakest



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Instantaneous dipoles are like ripples on the surface of a pool: On average, the surface of the pool is level, but the ripples make the water uneven. For an instant, there is more water in one part of the pool than in another. Similarly, the "ripples" caused by electron motion produce a slight attraction between neighboring molecules.

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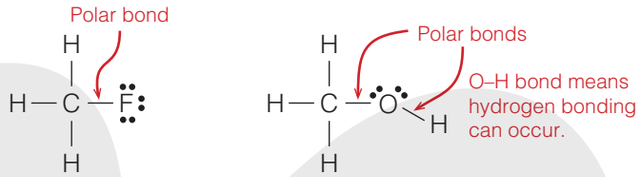
The stars are always in the sky, but their faint light is often blocked by the light of the moon or the bright light of the sun. In the same way, London dispersion forces are always present—but their effects are slight in molecules that have dipole–dipole or hydrogen-bonding interactions.

Example 10.1 Predicting the Properties of Compounds

Classify these three compounds, LiCl, CH_3F, and CH_3OH, as ionic, metallic, or covalent. Predict which compound would have the highest and lowest boiling points.

The first compound, LiCl, contains both a metal and a nonmetal. Because of this, we know that it is an ionic compound and therefore has high melting and boiling points.

The other compounds, CH_3F and CH_3OH, are composed entirely of nonmetal atoms and therefore are covalent (molecular) compounds. To predict the relative boiling points of these compounds, we need to draw Lewis structures and consider what polar bonds may be present:



Because F and O are much more electronegative than C or H, both of these molecules contain polar bonds and an overall dipole. However, CH_3OH contains an O–H bond. This means that CH_3OH molecules can form hydrogen bonds with each other, resulting in a higher boiling point.

Based on this analysis, we would expect the ionic compound, LiCl, to have the highest boiling point. Next highest is the compound that can form hydrogen bonds (CH_3OH), followed by the compound with only dipole–dipole interactions (CH_3F).

Check it
Watch Explanation

TRY IT

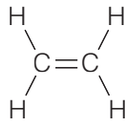
2. Classify the following as ionic, metallic, or covalent solids:

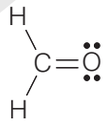
a. potassium nitrate **b.** phosphorus tribromide **c.** chromium

3. Which of these compounds would you expect to have the highest melting point, and why?

a. hexane, C_6H_{14} **b.** sodium fluoride, NaF

4. Based on their polarity, which of these compounds would you expect to have the highest boiling point?

a. 

b. 

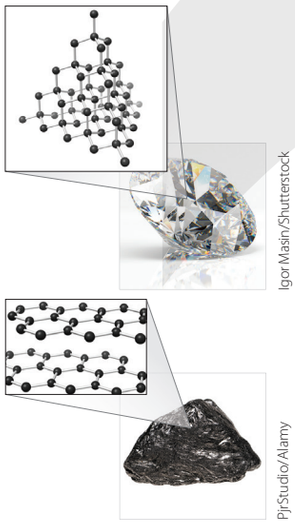


Figure 10.15 Diamond and graphite are both covalent networks of carbon. In diamond, each carbon has a tetrahedral geometry; in graphite, each carbon has a trigonal planar geometry.

Covalent Networks and Polymers

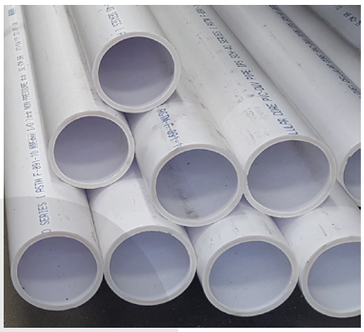
Some compounds contain covalent bonds but differ significantly from the molecular solids described earlier. For example, elemental carbon exists in two primary forms, *diamond* and *graphite*. Both of these are examples of **covalent networks**—lattices of connected covalent bonds, forming one giant molecule (**Figure 10.15**). In diamond, each carbon atom has four single bonds and a tetrahedral geometry. The single bonds repeat indefinitely in three dimensions,

resulting in a rigid structure that is among the hardest and most durable substances known. In graphite, the carbon atoms each have a trigonal planar geometry. The atoms are arranged in two-dimensional "sheets" that slide easily over each other. As a result, graphite is softer and is used as the writing material in most pencils.

Polymers are compounds that contain long chains of covalently bonded atoms. There are many naturally occurring polymers, including cellulose (the structural component of wood), starch (a major component of many staple foods, such as rice and potatoes), and proteins. *Plastics* such as poly(ethylene) and poly(vinyl chloride), or PVC, are synthetic polymers (**Figure 10.16**).

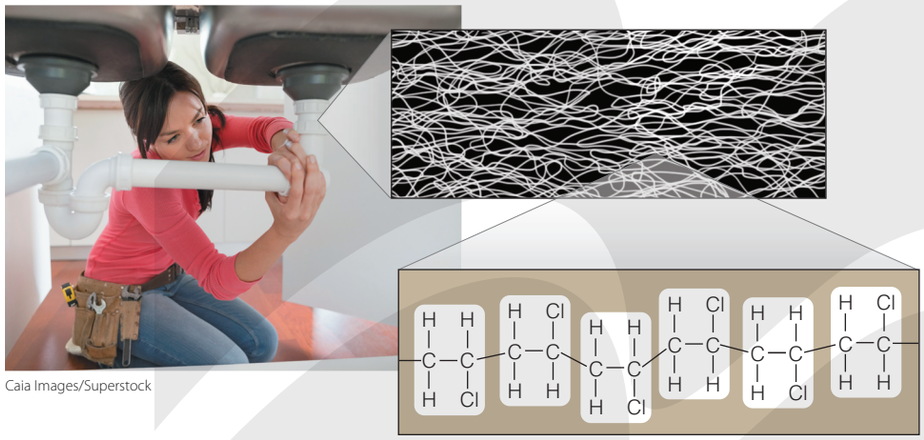
Polymers are composed of small, repeating units that are bonded together. For example, **Figure 10.17** shows the structure of poly(vinyl chloride). Each small unit is a single link in the longer chain. We'll discuss polymers in more detail in Chapter 15.

a_v_d/Shutterstock



Courtesy of Kevin Revell

Figure 10.16 Polyethylene is commonly used in disposable plastics like shrink wrap. Poly(vinyl chloride), or PVC, is commonly used in plumbing applications.



Caia Images/Superstock

Figure 10.17 Polymers like PVC are composed of long chains of small, repeating units. Each "link" of a PVC chain has the formula C_2H_3Cl.

Besides being used as a fuel source, fossil fuels provide the starting materials to make most plastics.

Plastics or polymers are made up of long chains of covalent bonds. ■

10.3 Describing Gases

In the previous section, we saw how the properties of solids and liquids depend on the forces of attraction between particles. In the remaining sections of this chapter, we'll shift our attention to gases. In a gas, the particles are spaced far apart. They move freely, interacting very little with the particles around them. Gas particles travel in a straight line until they bounce off another particle or off the walls of the container (**Figure 10.18**).

When describing gases, we often assume that they are behaving as *ideal gases*. An **ideal gas** has two key properties:

1. The volume of the particles is much, much less than the volume of the container.
2. The particles have no attraction for each other. When they pass each other, they do not slow down. When they collide, they bounce off each other like billiard balls.

When gases behave this way, we can predict their behavior mathematically by using simple relationships between the temperature of the gas, the volume the gas occupies, and the pressure it exerts.

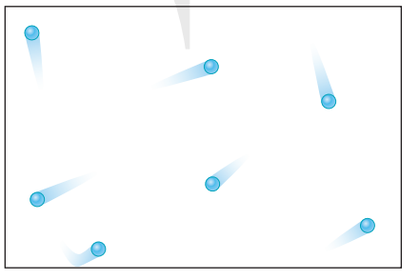


Figure 10.18 In a gas, particles move freely and have very little interaction with each other.

Explore
Figure 10.18

We commonly describe gases by the pressure they exert. ■

Pressure

We commonly describe gases by the **pressure** they exert on their surroundings. We can think of pressure as the "push" that particles exert on everything around them. For example, consider the gas represented in Figure 10.18. The particles move freely around the box. As they move, they collide with the walls of the container. Collectively, these tiny collisions create a larger push. This is the pressure on the container (**Figure 10.19**).

Formally, pressure is defined as the force that is exerted divided by the area over which it is applied:

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

As we will see in the following sections, the pressure caused by the gas depends on the volume, temperature, and amount of gas present.

Cathyrose Melloan/Alamy

Figure 10.19 The gas inside this rural propane tank exerts pressure on its walls.

Measuring Pressure

Gas pressure is a normal part of our lives. Earth's atmosphere exerts a pressure that is essential to life. Subtle variations in atmospheric pressure cause wind and weather patterns (**Figure 10.20**). Pressurized air in tires keeps vehicles rolling smoothly. The bounce of a basketball is caused by the compressed air inside it. Air-conditioning systems rely on the compression and expansion of gases under different pressures. Pneumatic tools use compressed air to produce a powerful force (**Figure 10.21**).

www.dennisoswald.de/Getty images

Figure 10.20 A tornado occurs when areas of high and low air pressure come together.

Photononstop/Alamy

Carolyn Franks/Alamy

Mark Hunt/Design Pics/Media Bakery

Figure 10.21 From tires to air-conditioning systems and pneumatic tools, compressed gas serves many useful purposes.

A **barometer** is a device used to measure atmospheric pressure. A classic barometer is shown in **Figure 10.22**. In this device, a long tube is filled with a liquid (usually mercury), then turned upside down in a reservoir of the liquid. Gravity pulls the liquid down, creating a vacuum in the top of the tube. At the same time, pressure from the outside air pushes the liquid up into the tube. The higher the air pressure, the higher the liquid is pushed into the tube. We measure the pressure of the atmosphere by measuring the height of the mercury inside the tube. Because of this, atmospheric pressure is often reported in **millimeters of mercury (mm Hg**; also called *torr*). At sea level, the average atmospheric pressure is 760 mm Hg, or 760 torr. This is referred to as *standard pressure*.

Why do barometers use mercury rather than a less-toxic liquid like water or oil? Because mercury is over 13 times denser than water. Atmospheric pressure is enough to push a column of mercury 760 millimeters (0.76 meters) into the air. If we used water in a barometer, the column would have to be over 33 feet (10 meters) high. Of course, there are many other designs for pressure gauges. The shapes and styles of these gauges differ based on their application (**Figure 10.23**).

For practical applications, we sometimes refer to the **gauge pressure** of a compressed gas. The gauge pressure is the difference between the compressed gas pressure and the atmospheric pressure. For example, if the gauge on an air compressor reads zero, it doesn't mean there is a vacuum (meaning no air) inside the cylinder; it means that the air pressure inside the cylinder is equal to the air outside. For the problems in this chapter, you can assume that the pressures given are absolute pressures, not gauge pressures.

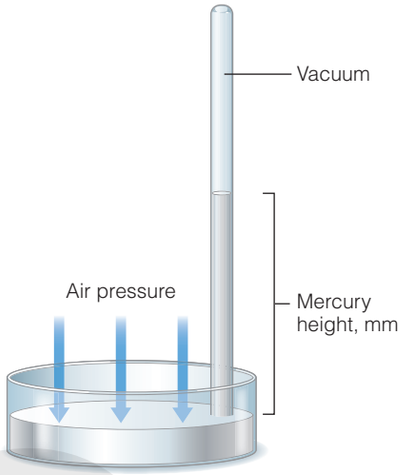


Figure 10.22 A simple barometer is used to measure the pressure of the atmosphere.

Explore Figure 10.22

1 torr = 1 mm Hg

Figure 10.23 Other uses for pressure gauges include (a) for measuring blood pressure, (b) for an acetylene tank used in welding, (c) for a scuba diving regulator, and (d) for measuring tire pressure. (a) Fotosr52/Shutterstock; (b) Sergiy Zavgorodny/Shutterstock; (c) Dmitry Kalinovsky/Shutterstock; (d) Science Source

When you pull the air out of a straw, you lower the air pressure inside the straw. Because the pressure outside is higher than inside, the liquid gets pushed up into the straw. A barometer works the same way.

TABLE 10.7 Common Pressure Conversions

1 atmosphere = 760 mm Hg (torr)
1 atmosphere = 101.3 kPa
1 atmosphere = 14.70 psi
1 atmosphere = 1.013 bar

We can use many different units to describe gas pressure. A common unit is the **atmosphere (atm)**, based on the standard air pressure at sea level. One atmosphere is defined as 760 mm Hg (29.92 inches Hg). In the United States, pressure is often described in terms of *pounds per square inch* (psi). A basketball has a gauge pressure of about 8 psi, while the tires on your car likely have a gauge pressure of 30–40 psi. Other common units of pressure include the *kilopascal* (kPa) and the *bar*. The relationships between these units are summarized in **Table 10.7**.

Check it
Watch Explanation

TRY IT

5. A tire has a maximum recommended pressure of 276 kPa. What is this pressure in psi?

6. Use the relationships given in Table 10.7 to complete these conversions:

a. Convert 2.4 atmospheres to psi.

b. Convert 0.892 kPa to atmospheres.

c. Convert 1,500 kPa to psi.

10.4 The Gas Laws

A scientific law describes an observed behavior, but it doesn't explain why the behavior occurs. The gas laws are an example of this idea.

The pressure, volume, and temperature of an ideal gas are related to each other through simple mathematical relationships called the **gas laws**. These relationships are very useful for describing common gases including air, nitrogen, oxygen, helium, acetylene, and carbon dioxide.

Boyle's Law

Boyle's law describes the relationship between the pressure and volume of a gas. According to Boyle's law, the pressure and the volume of a gas are inversely related. That is, if the volume (V) goes up, the pressure (P) goes down, and vice versa (**Figure 10.24**). As long as the temperature does not change, *the pressure times the volume of a gas is constant.*

$$PV = \text{constant}$$

Boyle's law is useful when the pressure or the volume is changing, and we wish to determine how the other unit will change. Because PV is constant, we can say that

$$P_1V_1 = P_2V_2$$

where P_1 and V_1 are the initial pressure and volume, and P_2 and V_2 are the final pressure and volume.

When working with Boyle's law, we can use any units of pressure and any units of volume. Example 10.2 describes this type of problem.

Wm. Baker/GhostWorx images/Alamy

Figure 10.24 A firefighter carries an air tank on his back. The air is compressed into a small volume, and so it has a high pressure.

Example 10.2 Using Boyle's Law

A commercial compressor stores 2.8 liters of air at a pressure of 150 psi. If this air is allowed to expand until the pressure is equal to 15 psi (just over atmospheric pressure), what volume will the air occupy?

In this example, we are looking for the volume of air after the pressure is decreased. P_1 is 150 psi, P_2 is 15 psi. V_1 is 2.8 L, and we are trying to solve for V_2. To solve this

problem, we rewrite the equation above to isolate V_2 and then plug in the pressure and volume quantities to solve:

$$V_2 = \frac{P_1V_1}{P_2} = \frac{(150\ \cancel{psi})(2.8L)}{15\ \cancel{psi}} = 28\ L$$

Notice that when we solved this problem, our units of pressure canceled out. It is important to make sure that the units of P_1 and P_2 are the same. The units for V_1 and V_2 will also be the same.

TRY IT

7. A balloon has a volume of 2.5 liters at a pressure of 1.0 bar. If the pressure around the balloon is decreased to 0.80 bar, what is the new volume of the balloon?

Check it
Watch Explanation

Charles's Law

Charles's law states that at constant pressure, the volume of a gas is directly proportional to its temperature. If the temperature goes up, the volume goes up. If the temperature goes down, the volume goes down. Mathematically, we can represent this by the following equation:

$$V \propto T$$

where the ∝ symbol means "is proportional to." Alternatively, we can write this as follows:

$$\frac{V}{T} = \text{constant}$$

Using Charles's Law to Find Absolute Zero

In the mid-nineteenth century, William Thomson (who was later granted the title Lord Kelvin) compared the plots of volume versus temperature for a number of different gas samples, similar to the graph in **Figure 10.25**. Although he could not measure gas pressures at very low temperatures, Thomson observed that if the volume continued to decrease in a linear fashion, all of the gases would reach a volume of zero at the same temperature. This temperature, –273.15 °C, is **absolute zero**—the lowest possible

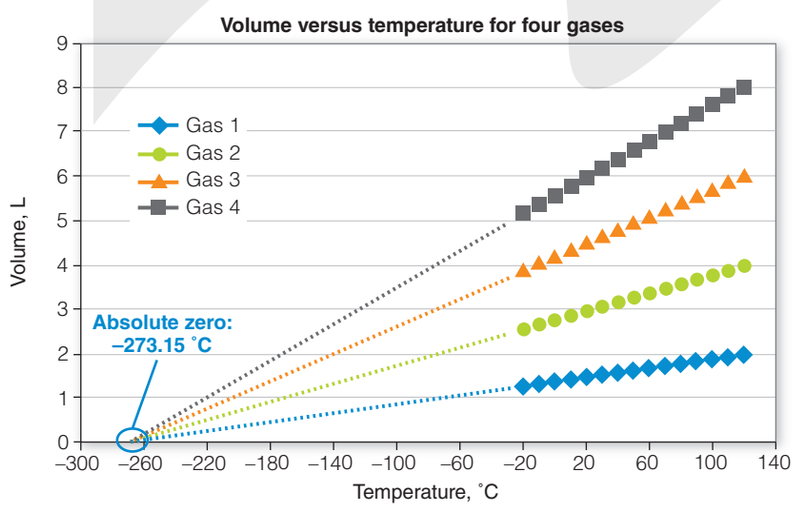


Figure 10.25 Absolute zero is calculated by extrapolating the volume–temperature relationship of a gas back to a volume of zero.

© Gerald D. Tang/TangsPhoto Stock

temperature. Recall that temperature is the measure of the kinetic energy of the particles in a substance. At absolute zero, the particles in a material have zero kinetic energy.

In the Kelvin temperature scale, absolute zero is given the value of 0 kelvin. A unit kelvin is the same size as a degree Celsius. We convert between the Celsius and Kelvin scales by using the following relationship:

$$\text{kelvin} = °\text{C} + 273.15$$

Solving Volume–Temperature Problems Using Charles's Law

We commonly use Charles's law to predict how the volume of a gas will change if the temperature changes, or vice versa. Because the ratio of volume to temperature is constant, we can write Charles's law using this relationship:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

where V_1 and T_1 are the initial volume and temperature, and V_2 and T_2 are the final volume and temperature.

When working a Charles's law problem, you can use any units of volume; but *you must express the temperature in kelvins*. Use of Celsius in a gas-law problem will give an incorrect answer. Example 10.3 illustrates this type of problem.

For gas-law problems, temperatures must be in kelvins.

Example 10.3 Using Charles's Law

A balloon has a volume of 3.2 liters at room temperature (25 °C). The gas inside the balloon is then heated to 100 °C. What is the new volume of the balloon?

In this example, we are looking for the volume of the balloon after the temperature has increased. V_1 is 3.2 L. Before solving for V_2, we must convert the temperatures to the Kelvin scale:

$$T_1 = 25\ °\text{C} + 273 = 298\ \text{K}$$

$$T_2 = 100\ °\text{C} + 273 = 373\ \text{K}$$

In these calculations, notice that we rounded 273.15 to 273 because we knew the Celsius temperature only to the nearest degree. Finally, we can rewrite the equation to isolate V_2, then plug in the quantities to solve.

$$V_2 = \frac{V_1T_2}{T_1} = \frac{(3.2\ \text{L})(373\ \cancel{\text{K}})}{298\ \cancel{\text{K}}} = 4.0\ \text{L}$$

Check it
Watch Explanation

TRY IT

8. A gas occupies a volume of 15.2 L at 25 °C. At what temperature would this gas expand to a volume of 30.4 L? Express your answer in both kelvins and degrees Celsius.

The Combined Gas Law

In many situations, the pressure, volume, and temperature of a gas all change together. For example, air-conditioning systems rely on the compression and expansion of gases. As the pressure and volume of the gas change, the temperature also changes. For situations like this, we use a mathematical relationship called the **combined gas law**:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

As in Boyle's and Charles's laws, we can use any units of pressure and volume, but the temperatures must be expressed in kelvins. Examples 10.4 and 10.5 illustrate the use of the combined gas law.

The combined gas law can be used to solve any Boyle's law or Charles's law problem. ■

Example 10.4 Using the Combined Gas Law

A gas with a temperature of 280 K, a pressure of 200 kPa, and a volume of 25.8 L is compressed to 15.8 L, causing the pressure to increase to 350 kPa. What is the temperature of the gas under the new conditions?

In this example, we're given P_1, V_1, and T_1, as well as P_2 and V_2. We're asked to solve for T_2. We can rearrange the combined gas law equation and insert the values to reach the answer.

$$T_2 = \frac{P_2V_2T_1}{P_1V_1} = \frac{(350\ \cancel{\text{kPa}})(15.8\ \cancel{\text{L}})(280\ \text{K})}{(200\ \cancel{\text{kPa}})(25.8\ \cancel{\text{L}})} = 300\ \text{K}$$

Example 10.5 Using the Combined Gas Law

Under constant-pressure conditions, a sample of hydrogen gas initially at 88 °C and 1.62 L is cooled until its final volume is 942 mL. What is its final temperature?

In this example, the pressure is constant, so $P_1 = P_2$. Because of this, it may be canceled from the equation. Also, notice that V_1 is given in liters, but V_2 is in milliliters: For the volume units to cancel, they must be the same. Therefore, let's express V_2 as 0.942 L. Converting the temperature to kelvins and substituting the values into the equation, we obtain a final answer of 210 K.

$$T_2 = \frac{\cancel{P_2}V_2T_1}{\cancel{P_1}V_1} = \frac{(0.942\ \cancel{\text{L}})(361\ \text{K})}{(1.62\ \cancel{\text{L}})} = 210\ \text{K}$$

TRY IT

9. On a cold Iowa morning when the temperature is –30 °C, a truck tire is inflated to a pressure of 45.0 psi. The truck is then driven south until the temperature reaches +33 °C. If the tire has not lost any air, what is the pressure in the tire after it warms up?

Check it
Watch Explanation

Avogadro's Law

Avogadro's law states that, if pressure and temperature are constant, the volume of a gas is proportional to the number of moles of gas present. The more gas that is present, the larger the volume it occupies. Mathematically, we can say that

$$V \propto n$$

where n is the number of moles of the gas.

The exact volume that a mole of gas occupies depends on its temperature and pressure. A temperature of 0 °C (273 K) and a pressure of 1.0 atmosphere is called *standard temperature and pressure*, or STP. At STP, one mole of gas occupies a volume of 22.4 liters (**Figure 10.26**).

Saturated/Getty images

Figure 10.26 At standard temperature and pressure (STP), 1 mole of gas occupies 22.4 L, the size of a large balloon.

Check it
Watch Explanation

TRY IT

10. At STP, 1 mole of gas occupies a volume of 22.4 L. At the same temperature and pressure, what volume does 5 moles of gas occupy?

The Ideal Gas Law

The gas laws described so far allow us to relate different properties of gases. We can connect all of these together using the **ideal gas law**. This extraordinarily useful law relates the amount of gas present to its pressure, volume, and temperature. We usually write this law as

$$PV = nRT$$

where P is the pressure, V is the volume, T is the temperature, n is the number of moles of gas, and R is a constant, called the *gas constant*, with a value of 0.0821 L·atm/mol·K.

When using the ideal gas law, it is important to make sure that the temperatures are expressed in kelvins and that the units of volume and pressure align with the values included in the gas constant (for example, liters and atmospheres). Examples 10.6 and 10.7 illustrate this law.

Remember, in an ideal gas, the particles do not interact with each other at all. Real gases will have some interactions, but most common gases behave in a nearly ideal manner.

Example 10.6 Using the Ideal Gas Law

What volume does 1.00 mole of gas occupy at a temperature of 0.00 °C and a pressure of 1.00 atmosphere?

To solve this problem, we rearrange the ideal gas law equation to isolate V and then insert the values:

$$V = \frac{nRT}{P} = \frac{(100\ \cancel{\text{mol}})\left(0.0821\dfrac{\text{L}\cdot\cancel{\text{atm}}}{\cancel{\text{mol}}\cdot\cancel{\text{K}}}\right)(273.15\ \cancel{\text{K}})}{(1.00\ \cancel{\text{atm}})} = 22.4\ \text{L}$$

Notice that in the solution above, all of the units cancel except for liters, and so the answer is given in liters. This problem corresponds to standard temperature and pressure (STP).

Example 10.7 Using the Ideal Gas Law

A portable oxygen tank has a volume of 2.40 L and a pressure of 243 psi at a temperature of 22 °C. How many moles of oxygen are present in this cylinder? What is the mass of the oxygen in grams?

To solve this problem, we first need to convert temperature to the Kelvin scale (22 °C = 295 K). To align our units with the gas constant, we also need to convert the pressure (243 psi = 16.5 atm). Using the ideal gas equation, we can then find the number of moles:

$$n = \frac{PV}{RT} = \frac{(16.5\ \cancel{\text{atm}})(2.40\ \cancel{\text{L}})}{\left(0.0821\dfrac{\cancel{\text{L}}\cdot\cancel{\text{atm}}}{\text{mol}\cdot\cancel{\text{K}}}\right)(295\ \cancel{\text{K}})} = 1.64\ \text{mol}$$

Once we know the moles, we can then convert to grams of oxygen, as described in Chapter 7.

$$1.64\ \cancel{\text{mol}} \times \frac{32.00\ \text{g}\ O_2}{1\ \cancel{\text{mol}}} = 52.5\ \text{g}\ O_2$$

TRY IT

11. At what temperature does 1.20 moles of hydrogen gas occupy a volume of 28.1 L at a pressure of 121 kilopascals?

12. A room with a volume of 50 m^3 has a pressure of 750 torr at a temperature of 72 °F. How many moles of gas occupy the room?

Check it
Watch Explanation

Mixtures of Gases

We commonly encounter gases that contain several components. For example, air is a mixture of gases, composed of about 78% nitrogen and 21% oxygen. The remaining 1% is a mixture of argon, carbon dioxide, water vapor, and other trace components (**Figure 10.27**).

Philip and Karen Smith/Getty Images

Figure 10.27 Air is a mixture of nitrogen, oxygen, argon, carbon dioxide, and smaller amounts of other gases.

Fortunately, the pressure and volume depend only on the amount of gas present, not on the identity of the gas. One mole of oxygen gas has the same pressure and volume as one mole of nitrogen gas or one mole of carbon dioxide gas.

When working with mixtures of gases, we sometimes use the **partial pressure** of the gases that are present. The partial pressure is the pressure caused by one gas in a mixture. We can add up the partial pressures to find the total pressure of the mixture. Example 10.8 illustrates this idea.

Adding up all the partial pressures gives the total pressure. ■

Example 10.8 Using Partial Pressures for a Mixture of Gases

If a 40.0-L cylinder is filled with 5.00 moles of nitrogen, 2.00 moles of oxygen, and 3.00 moles of carbon dioxide at a temperature of 400 K, what is the pressure inside the cylinder?

We could take two approaches in solving this problem. The first is to calculate the pressure that is produced by each of the individual gases. The partial pressures of nitrogen, oxygen, and carbon dioxide (labeled as P_{N_2}, P_{O_2}, and P_{CO_2}) are calculated as shown here:

$$P_{N_2} = \frac{nRT}{V} = \frac{(5.00\ \cancel{\text{mol}})\left(0.0821\dfrac{\cancel{\text{L}}\cdot\text{atm}}{\cancel{\text{mol}}\cdot\cancel{\text{K}}}\right)(400\ \cancel{\text{K}})}{(40.0\ \cancel{\text{L}})} = 4.11\ \text{atm}$$

$$P_{O_2} = \frac{nRT}{V} = \frac{(2.00\ \cancel{\text{mol}})\left(0.0821\dfrac{\cancel{\text{L}}\cdot\text{atm}}{\cancel{\text{mol}}\cdot\cancel{\text{K}}}\right)(400\ \cancel{\text{K}})}{(40.0\ \cancel{\text{L}})} = 1.64\ \text{atm}$$

$$P_{CO_2} = \frac{nRT}{V} = \frac{(3.00\ \cancel{\text{mol}})\left(0.0821\dfrac{\cancel{\text{L}}\cdot\text{atm}}{\cancel{\text{mol}}\cdot\cancel{\text{K}}}\right)(400\ \cancel{\text{K}})}{(40.0\ \cancel{\text{L}})} = 2.46\ \text{atm}$$

Once we've done this, we can find the total pressure by adding together the partial pressures:

$$P_{total} = P_{N_2} + P_{O_2} + P_{CO_2} = 4.11\ \text{atm} + 1.64\ \text{atm} + 2.46\ \text{atm} = 8.21\ \text{atm}$$

The second approach to this problem simplifies the calculation. Because the total pressure depends on the number of moles of gas (not the identity of the gas), we can find the total number of moles present and then calculate the pressure based on this combined value.

$$5.00\ \text{moles}\ N_2 + 2.00\ \text{moles}\ O_2 + 3.00\ \text{moles}\ CO_2 = 10.00\ \text{moles total}$$

$$P_{total} = \frac{nRT}{V} = \frac{(10.00\ \cancel{\text{mol}})\left(0.0821\dfrac{\cancel{\text{L}}\cdot\text{atm}}{\cancel{\text{mol}}\cdot\cancel{\text{K}}}\right)(400\ \cancel{\text{K}})}{(40.0\ \cancel{\text{L}})} = 8.21\ \text{atm}$$

Check it
Watch explanation

TRY IT

13. An air sample contains 0.3% water vapor. If the total air pressure is 765 torr, what is the partial pressure due to the water?

14. A 24-liter cylinder contains 150.0 grams of NH_3 gas and 600.0 grams of argon gas at a temperature of 300 K. What is the pressure inside the cylinder?

15. While walking along the ocean, you take a deep breath. If your lung capacity is 5.0 liters, the temperature is 87 °F, and the pressure is 765 torr, how many moles of air can you take in? Since air is composed of about 21% oxygen (by mole percent), how many grams of oxygen have you taken in?

A Molecular View of the Gas Laws

Pressure increases with temperature. ■

In the preceding sections, we used mathematical relationships to describe how gases behave. By thinking about gases on the atomic or molecular level, we can begin to understand why these mathematical relationships are so, and we can also predict the behavior of gases in other situations. Let's ask three questions about the relationship between the particles in a container and the pressure they exert:

1. *How does the pressure change if the temperature of the gas increases?* At higher temperatures, the particles move faster. Faster-moving particles strike their surroundings with more force, exerting more pressure. On the other hand, if the temperature decreases, the pressure drops. This is why tires are sometimes partially flat on cold mornings (**Figure 10.28**).
2. *What happens to the pressure of a gas if the volume increases?* If the volume of a container increases, the particles have to travel farther before reaching the walls of the container. As a result, the particles collide with the container less frequently. An increase in volume produces a decrease in pressure.
3. *What happens to the pressure of a gas if the number of particles in a container is increased?* If you've ever pumped up a tire, you know that increasing the amount of gas present increases the pressure. On a molecular level, more gas particles mean the particles will strike the sides of the container more often, creating more pressure. The dependence of pressure on temperature, number of particles, and volume of the container is shown in **Figure 10.29**.

Keith Beaty/Getty Images

Figure 10.28 What happens to the pressure in bike tires on a cold morning? Why is this so? What happens to the pressure if we pump more gas particles into the tire?

Increasing the amount of gas increases the pressure. ■

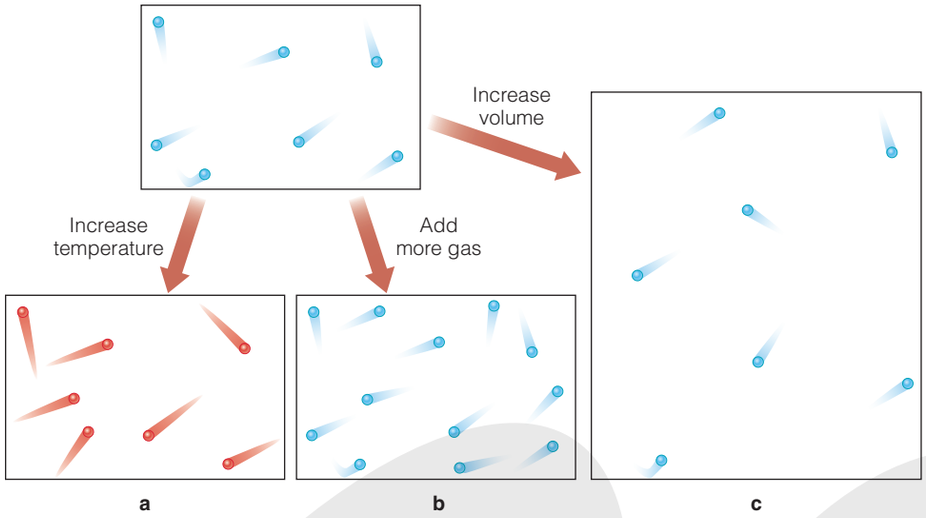


Figure 10.29 How does the pressure change as the temperature, volume, and number of particles change? (a) Increasing temperature increases pressure. (b) Increasing the number of particles increases the pressure. (c) Increasing the volume decreases the pressure.

Explore
Figure 10.29

TRY IT

16. How do these changes affect the pressure of a gas?

a. decreasing the temperature

b. decreasing the volume

c. decreasing the amount of gas

Check it
Watch Explanation

10.5 Diffusion and Effusion

Diffusion is the spread of particles through random motion. Diffusion can refer to either the liquid or the gaseous state, but the principle is the same: Particles randomly move, and as they do, they slowly spread from areas of higher concentration to lower concentration. For example, gases emitted from a smokestack diffuse out into the atmosphere (**Figure 10.30**). Because lighter particles move more quickly than heavier particles, lighter gases diffuse more quickly than heavier gases.

Effusion is the process of a gas escaping from a container. Like diffusion, effusion depends on the velocity of the gas particles. For example, **Figure 10.31** shows two containers, one filled with helium, one filled with air. Each container has a single small hole out of which the gases can escape. The smaller helium particles move faster and therefore collide with the container walls more frequently. As a result, they are more likely to encounter the opening, and they leak out of the container more quickly.

Westend61/Superstock

Figure 10.30 Gases spread from areas of high concentration to low concentration. This process is called *diffusion*.

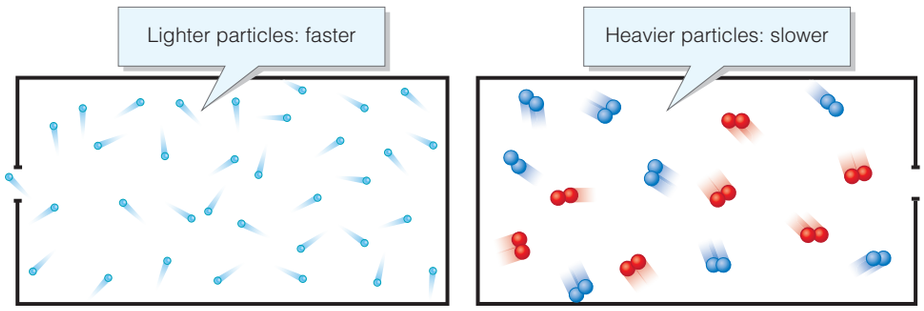


Figure 10.31 Because they move faster, lighter gases escape from a container more quickly.

Explore
Figure 10.31

Lighter gases diffuse and effuse faster. ■

We can see the principle of effusion in the way balloons lose air. If we fill one rubber balloon with helium, another with nitrogen, and another with argon, we'll see that the balloon filled with helium (the lightest gas) goes flat most quickly, while the balloon filled with argon (the heaviest gas) goes flat most slowly (**Figure 10.32**). For this reason, helium balloons are typically made of heavier, less porous materials.

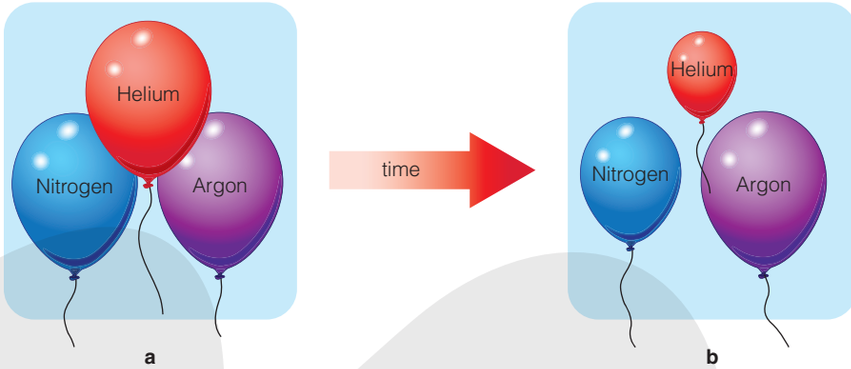


Figure 10.32 (a) When balloons are filled with different gases, (b) the one containing the lightest gas goes flat most quickly.

Check it
Watch Explanation

TRY IT

17. Four containers of gas are opened to the air at the same time. One contains carbon dioxide (CO_2), one contains methane (CH_4), one contains propane (C_3H_8), and one contains carbon disulfide (CS_2). Which of these gases would mix with the air most quickly? Which would mix most slowly?

10.6 Gas Stoichiometry

Sudden changes in gas pressure can have explosive results. These changes often arise when gases form or are consumed in a chemical reaction. To safely conduct reactions involving gases, we must know how much gas is produced or consumed. For example, the process of fermentation, used in making beer and bread (**Figure 10.33**), involves the reaction of glucose to form ethanol and carbon dioxide gas:

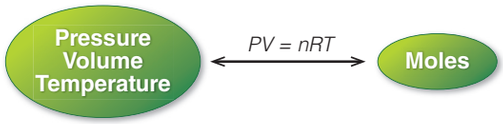
$$C_6H_{12}O_6\,(s) \rightarrow 2\,C_2H_6O\,(l) + 2\,CO_2\,(g)$$

In this fermentation reaction, how much carbon dioxide is produced for each kilogram of glucose that reacts?

To answer this question, we must combine the rules of stoichiometry (covered in Chapter 7) with the gas laws. At the center of these two concepts is the mole. Using the ideal gas law, we can relate P, V, and T to the number of moles of a gas:

Scott Eisen/Bloomberg via Getty images

Figure 10.33 These copper boiling and fermentation tanks are from the Samuel Adams Brewhouse in Boston, Massachusetts. Because the fermentation reaction produces carbon dioxide gas, brewers must monitor the pressure of gases inside the tanks.



In Chapter 7 we first talked about a "mole map" that showed the relationships between grams, moles, particles, and the balanced equation (see Figure 7.13). Now we can expand that map to include gases as well (**Figure 10.34**). Examples 10.9 and 10.10 demonstrate how to solve these problems.

The mole map, including gases

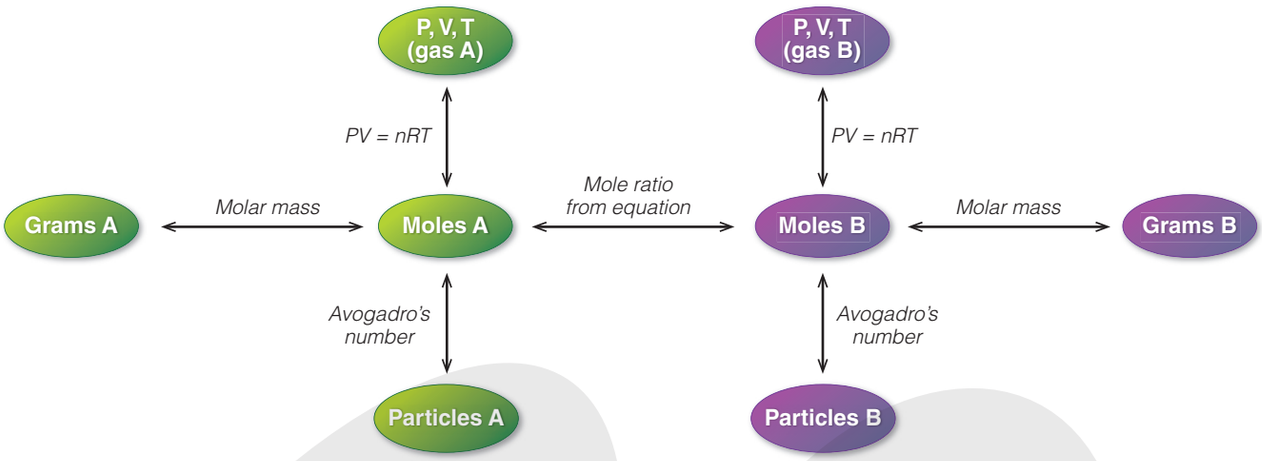


Figure 10.34 The mole is the central hub for conversion between units.

Example 10.9 Gas Stoichiometry

In the fermentation of glucose, how many moles of carbon dioxide are produced for each kilogram of glucose that reacts? If the reaction takes place in a sealed container and the gas occupies a volume of 8.10 liters at a temperature of 21 °C, find the pressure of the carbon dioxide gas inside the container.

$$C_6H_{12}O_6\,(s) \rightarrow 2\,C_2H_6O\,(l) + 2\,CO_2\,(g)$$

To solve this problem, we first relate grams of glucose to moles of carbon dioxide:

$$\text{Grams } C_6H_{12}O_6 \rightarrow \text{Moles } C_6H_{12}O_6 \rightarrow \text{Moles } CO_2$$

We begin with 1 kg (1,000 g) of glucose:

$$1{,}000\ \cancel{\text{g } C_6H_{12}O_6} \times \frac{1\ \cancel{\text{mole } C_6H_{12}O_6}}{180.18\ \cancel{\text{g } C_6H_{12}O_6}} \times \frac{2\ \text{moles } CO_2}{1\ \cancel{\text{mole } C_6H_{12}O_6}} = 11.10\ \text{moles } CO_2$$

Once we've found the moles of CO_2, we use the ideal gas equation to find the pressure. Be sure to convert the temperature to kelvins and then solve:

$$P = \frac{nRT}{V} = \frac{(11.10\ \cancel{\text{mol}}\ CO_2)\left(0.0821\dfrac{\cancel{\text{L}}\cdot\text{atm}}{\cancel{\text{mol}}\cdot\cancel{\text{K}}}\right)(294\ \cancel{\text{K}})}{(8.10\ \cancel{\text{L}})} = 33.1\ \text{atm } CO_2$$

A pressure of 33.1 atm is a dangerous explosion risk. To prevent this risk, brewers must be careful to release some of the CO_2 formed and to monitor the pressure inside fermentation vessels.

In this example, we first used the balanced equation to solve for moles; then we used the ideal gas law to convert from moles to pressure. In other problems, we may be given the pressure and volume of a gas and asked to relate it to another reagent. In this case, we first convert to moles using the ideal gas law and then solve the stoichiometry problem. Example 10.10 illustrates this process.

Practice
Gas Stoichiometry
Try this interactive for additional practice with gas stoichiometry problems.

Vitaly Maselko/Alamy

Figure 10.35 The characteristic blue flame of a natural gas stove.

Example 10.10 Gas Stoichiometry

*Natural gas burns cleanly in air, according to the equation shown below (***Figure 10.35***):*

$$CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (g)$$

If 13.1 liters of CH_4 burn at a pressure of 1.00 atmosphere and a temperature of 290 K, what mass of carbon dioxide gas is produced?

In this question, we're given the pressure and volume of a gas. Our first step is to convert this to moles, using the ideal gas equation:

$$n = \frac{PV}{RT} = \frac{(1.00\ \cancel{\text{atm}})(13.1\ \cancel{\text{L}})}{\left(0.0821 \frac{\cancel{\text{L}} \cdot \cancel{\text{atm}}}{\text{mol} \cdot \cancel{\text{K}}}\right)(290\ \cancel{\text{K}})} = 0.550\ \text{mol}\ CH_4$$

Once we have the moles of CH_4, we can use the balanced equation to relate this to the moles of CO_2. We can then convert from moles of CO_2 to grams of CO_2, as shown here:

$$0.550\ \cancel{\text{moles}\ CH_4} \times \frac{1\ \cancel{\text{mole}\ CO_2}}{1\ \cancel{\text{mole}\ CH_4}} \times \frac{44.01\ \text{g}\ CO_2}{1\ \cancel{\text{mole}\ CO_2}} = 24.2\ \text{g}\ CO_2$$

Because these problems involve multiple steps, it will take some time before you become proficient at them. I encourage you to try the following problems.

Check it
Watch Explanation

TRY IT

18. In carbonated soft drinks, the carbon dioxide reacts with water to produce carbonic acid. This gives the drinks a pungent flavor that is balanced with a large amount of sugar. Over time, carbonic acid converts back into carbon dioxide and water, as shown in the equation below. What mass of H_2CO_3 would be required to produce a volume of 2.0 liters of CO_2 at a pressure of 1.0 atm and a temperature of 298 K?

$$H_2CO_3\ (aq) \rightarrow H_2O\ (l) + CO_2\ (g)$$

19. An airbag contains a solid fuel, which must react very rapidly to produce a large amount of gas. Airbags typically use several reactions in tandem, one of which is the decomposition of sodium azide to give sodium and nitrogen:

$$2\ NaN_3\ (s) \rightarrow 2\ Na\ (s) + 3\ N_2\ (g)$$

In this reaction, what volume of gas would be produced by the decomposition of 15 grams of NaN_3 if the temperature is 350 K and the pressure is 1.2 atm?

20. Acetylene torches burn at very high temperatures, using this reaction:

$$2\ C_2H_2\ (g) + 5\ O_2\ (g) \rightarrow 4\ CO_2\ (g) + 2\ H_2O\ (g)$$

At standard temperature and pressure, what volume of oxygen gas is required to completely react with a 4.8-liter cylinder of acetylene having an absolute pressure of 80 psi and a temperature of 20 °C?

Capstone Question

At the beginning of this chapter, we introduced the fossil fuels—coal, oil, and natural gas. Many fleet vehicles use compressed natural gas as an alternative to liquid gasoline (**Figure 10.36**). Natural gas is mainly composed of methane, CH_4. This equation shows the combustion of methane:

$$CH_4\,(g) + 2\,O_2\,(g) \rightarrow 2\,H_2O\,(g) + CO_2\,(g)$$

Using the table below, draw the Lewis structure for each substance in this reaction, and determine the electronic and molecular geometries for CH_4, CO_2, and H_2O. Compare the types of intermolecular forces in liquid CH_4 and liquid H_2O. Which substance has the stronger intermolecular forces? How does this impact their boiling points?

Methane burns with a fuel value of 55.5 kJ/g. If a 1.00-L cylinder containing pure methane gas at a pressure of 5,000 psi at 298 K is completely burned, how much energy will be released? What mass of CO_2 will form in this reaction?

Substance	Mass (g/mol)	Boiling Point	Lewis Structure	Electronic Geometry	Molecular Geometry
CH_4	16.0	–162 °C			
O_2	32.0	–183 °C			
CO_2	44.0	—*			
H_2O	18.0	100 °C			

*At standard atmospheric pressure, CO_2 sublimes (converts from solid to gas) at –79 °C.

Capstone Video

Gado Reportage/Alamy

Figure 10.36 Many fleet vehicles like this delivery truck run on compressed natural gas.

SUMMARY

In this chapter, we've explored the structure and properties of solids, liquids, and gases. In a solid, the particles pack closely together in fixed positions. In a liquid, the particles are close together but are not held in a fixed position. In a gas, the particles are far apart and have little or no interaction with the particles around them. The melting point and boiling point of a substance depend on how strongly the particles are held together.

The properties of a substance depend on its composition and structure. Ionic compounds form rigid lattices with each particle strongly attracted to those around it. Ionic compounds have very high melting and boiling points. In metallic substances, the atoms pack closely together and share electrons loosely between multiple atoms. This structure creates strong interactions, but not the rigid framework of ionic compounds.

Molecular compounds contain discrete groups of atoms, called molecules, that are connected by covalent bonds. The physical properties of molecular compounds depend on the interactions between molecules. Three major types of intermolecular forces are London dispersion forces, dipole–dipole interactions, and hydrogen bonds. Dispersion forces are the weakest interactions, while hydrogen bonds are the strongest. Molecular compounds typically have lower melting and boiling points than ionic compounds.

Some substances, such as diamond and graphite, form extensive networks of covalent bonds rather than small molecules. Polymers are compounds that contain long chains of covalent bonds. Plastics are synthetic polymers.

Gases have neither fixed shape nor fixed volume. In an ideal gas, the particles do not interact with each other and occupy a negligible fraction of the volume of the container.

Gases are typically measured by their pressure. A barometer is a device that measures atmospheric pressure. Pressure is measured in several different units, including atmospheres, torr (also called mm Hg), pounds per square inch (psi), bars, and kilopascals (kPa).

Boyle's law states that at constant temperatures, the product of the pressure and volume of an ideal gas are constant. This means that as pressure increases, the volume decreases, and vice versa.

Charles's law states that the volume of a gas is directly proportional to its temperature at constant pressure. By extrapolating the volume of a gas down to zero, we can identify the lowest possible temperature, called absolute zero. At this temperature, particles have no kinetic energy. The Kelvin temperature scale measures the absolute temperature of a substance: Absolute zero is given the value of 0 kelvin. When working with the gas laws, we must convert temperatures to the Kelvin scale.

The combined gas law integrates Charles's and Boyle's laws into a single law that relates the pressure, volume, and temperature of a gas. This law is most commonly expressed as

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

where P, V, and T denote the pressure, volume, and temperature, and the subscripts denote initial and final conditions.

Avogadro's law states that at constant pressure, the volume of a gas is proportional to the number of moles present. At standard temperature and pressure (STP), one mole of gas occupies a volume of 22.4 liters.

The ideal gas law describes the relationship of the P, V, and T to the number of moles of gas present. The relationship is given by the expression

$$PV = nRT$$

where n is the number of moles of gas, and the gas constant, R, has a value of 0.0821 L·atm/mol·K.

In a mixture of gases, each component contributes to the overall pressure of the system. The total pressure of the system is simply the sum of the partial pressures of the gases present.

We can understand the relationships described in the gas laws by thinking about gases on a molecular level. Because the pressure inside a container arises from the collisions of gases with the walls of the container, an increase in pressure occurs as the amount of gas increases (more particles striking the walls), as temperature increases (faster-moving particles strike the walls more frequently and with more energy), or as the volume decreases (the particles strike the walls more frequently).

Diffusion is the spread of particles through random motion, such as the motion of gas particles. Because lighter gas particles move more quickly than heavier particles, lighter gas particles diffuse more quickly. A related term, *effusion*, refers to the escape of gas from a container. Because smaller particles move faster, they also effuse more quickly than heavier particles.

As we saw in Chapter 7, we can use a balanced equation to predict the grams and moles that will be consumed or produced in a reaction. By combining stoichiometry calculations with the ideal gas law, we can predict changes in pressure and volume that accompany reactions involving gases.

Rethinking Gas Storage

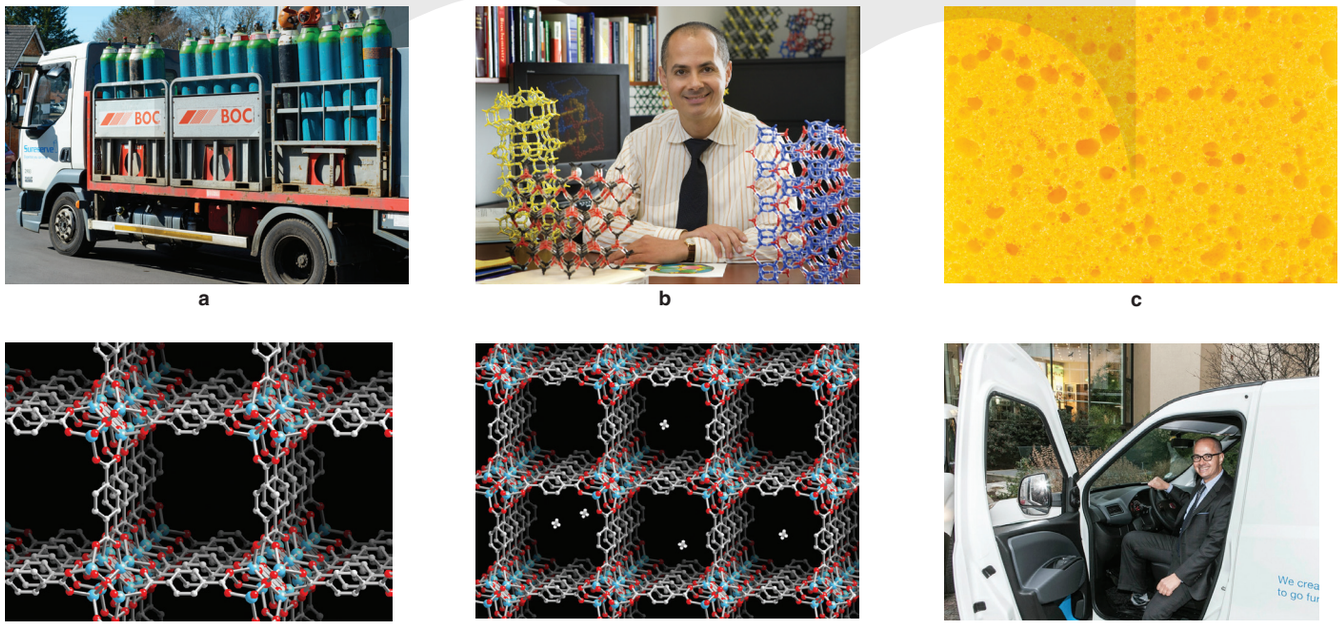
Compressed gases are a part of modern life. From fountain drinks to air conditioners and from mechanical work to health care, compressed gases are all around us. But compressed gases can be hard to handle. Because of their high pressures, gas cylinders require heavy steel walls. They can be difficult to transport, and they are potentially dangerous. If a gas cylinder ruptures, the pressure inside the container can fire the cylinder like a missile.

Omar Yaghi is a chemist at the University of California, Berkeley (**Figure 10.37**). For the past two decades, he has worked on a class of compounds called *metal-organic frameworks*, or MOFs, that hold the potential to change the way people store gases.

Originally from Jordan, Yaghi arrived in the United States when he was 15. Although he spoke very little English, he enrolled in classes at a community college near Albany, New York. His abilities and hard work paid off, and he eventually earned a Ph.D. in chemistry from the University of Illinois. As he describes it, "I was drawn to the beauty of chemistry. I didn't necessarily want to change the world—I was driven by curiosity and passion for science."

The MOF compounds that Yaghi creates *are* beautiful, but also very useful. Like a sponge or a honeycomb, these substances contain many tiny pores—except they are much smaller. Normally, gases move freely around a container, interacting very little with their surroundings. A MOF is able to trap gases within the tiny pores (Figure 10.37e). As a result, there are fewer free particles, and the pressure of the gas decreases. If a gas is placed in a cylinder that contains MOFs, the pressure is lower than if the gas were placed in an empty cylinder!

MOF technology has created a surge of interest in gas storage. For example, several companies are developing MOFs for the fuel tanks of vehicles powered by natural gas. Vehicles using this technology can safely store more fuel, so they have to be refueled less frequently.



a b c d e f

Figure 10.37 (a) The transport and storage of gas requires heavy steel containers. (b) Omar Yaghi is a chemist at UC Berkeley who studies gas storage. (c) A sponge is an example of a porous material. The empty spaces in these structures can trap other materials. (d) Yaghi's compounds, called MOFs, contain tiny pores. (e) These pores can trap gas particles, reducing the number of free particles and decreasing the pressure in the container. This image depicts CH_4 molecules trapped in a MOF. (f) Dr. Yaghi sits behind the wheel of a vehicle operating on a MOF methane fuel tank.
(a) DWD-photo/Alamy; (b) Omar M. Yaghi Research group at University of California Berkeley; (c) ©Ksena32/depositphotos.com; (d) Omar M. Yaghi Research group at University of California Berkeley; (e) Omar M. Yaghi Research group at University of California Berkeley; (f) Omar M. Yaghi Research group at University of California Berkeley

Key Terms

10.2 Solids and Liquids

intermolecular forces The forces of attraction or repulsion that take place between molecules.

dipole–dipole interaction An intermolecular force between two molecules containing net dipoles.

hydrogen bond An unusually strong dipole–dipole interaction that occurs between molecules containing H–F, H–O, or H–N bonds.

London dispersion forces Intermolecular forces that result from fluctuations in charge density called *instantaneous dipoles*.

covalent networks Long two- or three-dimensional sequences of covalent bonds, resulting in very large single molecules.

polymer A compound composed of long chains of covalently bonded atoms.

10.3 Describing Gases

ideal gas A gas in which the volume of the particles is much less than the volume of the container, and in which the particles have no attraction for each other.

pressure The force that an object exerts divided by the area over which it is applied; for gases, pressure describes the force that gases exert on their surroundings.

barometer A device used to measure atmospheric pressure.

millimeters of mercury (mm Hg) A measure of gas pressure; this unit originates from the height to which atmospheric pressure can push a column of mercury in a barometer; 1 mm Hg = 1 torr.

gauge pressure The difference between a compressed gas pressure and atmospheric pressure.

atmosphere (atm) A unit of gas pressure; 1 atm = 760 mm Hg.

10.4 The Gas Laws

gas laws Mathematical relationships between the pressure, volume, and temperature of gases.

Boyle's law The pressure and volume of an ideal gas are inversely related; the product of PV is constant at constant temperature.

Charles's law The volume of an ideal gas is directly proportional to its temperature; the relationship between V and T is constant at constant pressure.

absolute zero The lowest possible temperature, corresponding to 0 K or –273.15 °C; at this temperature, the particles in a substance have zero kinetic energy.

combined gas law A combination of Boyle's law and Charles's law; it states that for an ideal gas, the quantity PV/T is constant; usually expressed by the equation $P_1V_1/T_1 = P_2V_2/T_2$, where the subscripts 1 and 2 denote two different conditions.

Avogadro's law If pressure and temperature are constant, the volume of a gas is proportional to the number of moles of gas present.

ideal gas law The relationship between pressure, volume, temperature, and the number of moles of an ideal gas; typically expressed in the form $PV = nRT$, where R is the gas constant.

partial pressure The pressure caused by one gas in a mixture.

10.5 Diffusion and Effusion

diffusion The spread of particles through random motion; lighter gases diffuse more quickly than heavier gases.

effusion The process of a gas escaping from a container; lighter gases effuse more quickly than heavier gases.

Additional Problems

10.1 Interactions between Particles

21. Describe the arrangement and motion of particles in a solid, liquid, and gas.

22. How does the motion of particles change as a substance transitions from liquid to gas? In this transition, does the substance absorb or release heat energy?

10.2 Solids and Liquids

23. How is the arrangement of particles different between an ionic and a metallic solid?

24. Are there examples in which an ionic solid contains covalent bonds?

25. What are the *intra*molecular forces in a molecular solid? What types of *inter*molecular forces can exist in a covalent solid?

26. What types of covalent bonds are necessary for an intermolecular hydrogen bond to form?

27. Describe each of the following as ionic, metallic, or molecular solids:

a. calcium fluoride
b. glucose, $C_6H_{12}O_6$
c. bronze, an alloy of copper and tin
d. table salt, NaCl

28. Determine whether the following properties broadly describe ionic, metallic, or molecular solids. Some properties may describe more than one group.

a. high melting point
b. malleable
c. low boiling point
d. loosely shared electrons within a network of atoms

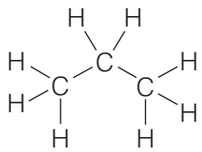
29. These four compounds have very similar formula masses. Classify them as ionic or covalent. Predict which of the compounds would have the highest and lowest boiling points.

a. LiF **b.** H_2O **c.** N_2 **d.** HCl

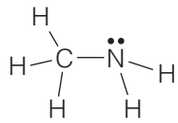
30. These four compounds have very similar formula masses. Classify them as ionic or covalent. Predict which of the compounds would have the highest and lowest boiling points.

a. Li_2S **b.** HF **c.** CH_3F **d.** O_2

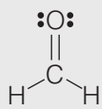
31. The Lewis structure of propane, C_3H_8, is shown here. What is the strongest type of intermolecular force in liquid propane?



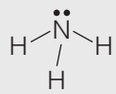
32. The Lewis structure of methylamine, CH_3NH_2, is shown here. What is the strongest type of intermolecular force in liquid methylamine?



33. The Lewis structure of formaldehyde is shown here. What is the strongest type of intermolecular force in liquid formaldehyde? Sketch two molecules, using a dashed line (as in Figure 10.6) to show the positive region of one molecule interacting with the negative region of another molecule.



34. The Lewis structure for ammonia, NH_3, is shown here. What is the strongest type of intermolecular force in liquid ammonia? Sketch two molecules, using a dashed line (as in Figure 10.6) to show the positive region of one molecule interacting with the negative region of another molecule.



35. Draw Lewis structures for each of the following molecules. Determine whether each molecule has a net dipole, and identify the strongest intermolecular force that would act between molecules of each pure substance.

a. CH_4 **b.** CH_2F_2 **c.** HNO (N is the central atom)

36. Draw Lewis structures for each of the following molecules. Determine whether each molecule has a net dipole, and identify the strongest intermolecular force that would act between molecules of each pure substance.

a. HCN **b.** HOCl (O is the central atom) **c.** CS_2

37. Methanol (CH_3OH) forms hydrogen bonds. Sketch Lewis structures for two methanol molecules, and use dashed lines to show the hydrogen bonds that can form.

38. Methanol (CH_3OH) dissolves easily in water because hydrogen bonds can form between these molecules. Sketch a methanol molecule and a water molecule, and show two possible hydrogen bonding interactions between the two mole cules.

39. What are polymers? How is a polymer different from a molecular solid?

40. How is a covalent network different from a molecular solid?

10.3 Describing Gases

41. Describe the motion of particles in a gas.

42. What are the two criteria for a gas to be considered an ideal gas?

43. Describe how the pressure on a container arises from the particles inside and outside the container.

44. What happens if the pressure inside a sealed container is much greater than the pressure outside the container? What happens if the pressure outside a container is much greater than the pressure inside?

45. When drinking through a straw, you are able to control the height of the liquid inside the straw by changing the pressure inside your mouth, as shown in **Figure 10.38**. What happens if the pressure in your mouth is lower than the air pressure outside? What happens if the pressure in your mouth is higher than the air pressure outside?

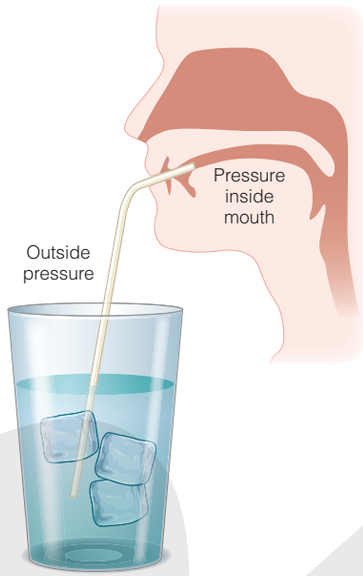
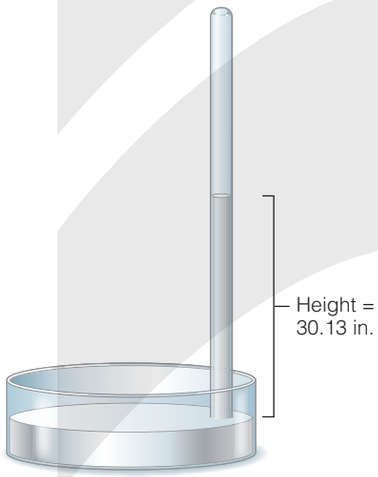


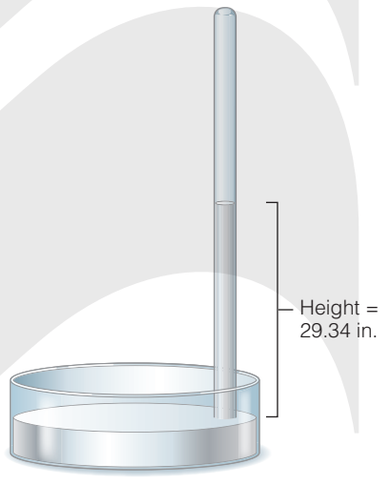
Figure 10.38 A straw is like a barometer.

46. Ignoring the fact that humans can't survive in a vacuum, is it possible to drink through a straw in the vacuum of space? Why or why not?

47. Measure the pressure shown on the barometer. Is this pressure higher or lower than standard atmospheric pressure? (1 inch = 25.4 mm)



48. Measure the pressure shown on the barometer. Is this pressure higher or lower than standard atmospheric pressure? (1 inch = 25.4 mm)



49. In August 1992, Hurricane Andrew slammed into the Miami area, causing 23 deaths and destroying over 25,000 homes. When it made landfall, the pressure in the eye of the storm was 922 millibars. What is this pressure in torr? Is this pressure higher or lower than standard atmospheric pressure?

50. In November 2013, Typhoon Haiyan struck the Philippine islands, causing over 6,000 deaths. At one point, the pressure in the eye of the storm was measured at 26.43 inches of mercury. What is this pressure in torr? What is this pressure in millibars? Is this pressure higher or lower than standard atmospheric pressure?

51. Express standard atmospheric pressure in millibars, torr, and inches of mercury.

52. Express standard atmospheric pressure in mm Hg, kilopascals, and bars.

53. Convert the following pressures:

a. Express 698 mm Hg in millibars.
b. Express 3.2 atm in torr.
c. Express 1.42 bars in psi.

54. Convert the following pressures:

a. Express 1.2 atmospheres in bars and millibars.
b. Express 32.41 psi in torr.
c. Express 23.29 inches of mercury in atmospheres.

55. Gauge pressure refers to the difference in pressure between a compressed gas and atmospheric pressure. If the atmospheric pressure is 15.0 psi, and you inflate your bicycle tire to a gauge pressure of 110.0 psi, what is the absolute pressure in the tires?

56. Gauge pressure refers to the difference in pressure between a compressed gas and atmospheric pressure. If the atmospheric pressure is 15.0 psi, and you inflate your football to a gauge pressure of 10.0 psi, what is the absolute pressure in the football?

10.4 The Gas Laws

57. How does the pressure change if the volume of a gas decreases? How does it change if the temperature decreases?

58. If the volume of a gas increases, how does the pressure change? How does the pressure change if the temperature of a gas increases?

59. A gas with an initial pressure of 1.0 atmosphere and a volume of 4.3 liters is compressed to a pressure of 5.1 atmospheres. What is the new volume of the gas?

60. A gas with an initial pressure of 780 torr and a volume of 150 liters is compressed to a volume of 32 liters. What is the pressure of the compressed gas?

61. A balloon with a pressure of 15.0 psi has a volume of 1.24 liters. If the pressure drops, the balloon will expand. What will be the volume of the balloon if the pressure decreases to 10.4 psi?

62. A pressurized gas is allowed to expand. If the gas originally occupied a volume of 420 mL at a pressure of 300 psi, what volume will it occupy if pressure is lowered to 150 psi?

63. A balloon with a volume of 1.41 liters at a temperature of 300.0 K is heated to 350.0 K. What is the new volume of the balloon?

64. A balloon with a volume of 4,320 cm^3 at a temperature of 25 °C is heated to 95 °C. What is the new volume of the balloon?

65. An ideal gas is allowed to cool at a constant pressure. The gas occupies a volume of 20.0 L at a temperature of 273.15 K (0 °C). At what temperature will it occupy a volume of 10.0 L?

66. A gas occupies a volume of 800.0 mL at a temperature of 25 °C. At what temperature would the gas occupy only half this volume?

67. What is absolute zero? Describe absolute zero with regard to the motion and kinetic energy of molecules.

68. Is it possible for a substance to be a gas at absolute zero?

69. A gas has a pressure of 900.0 millibars at a temperature of 30.0 °C. If the volume is unchanged but the temperature is increased to 80.0 °C, what is the new pressure of the gas?

70. A gas occupying a volume of 43.0 liters has a pressure of 850 millibars at a temperature of 35 °C. If the volume is unchanged, but the temperature is increased to 120 °C, what is the new pressure of the gas?

71. A gas cylinder at a temperature of 0 °C has a pressure of 22.4 psi. At what temperature would the pressure inside the cylinder increase to 35.0 psi?

72. A tank of gas at a temperature of 0 °C has a pressure of 415 psi. At what temperature would the pressure inside the tank increase to 600 psi?

73. A gas inside a piston initially has a pressure of 1.2 bars at a temperature of 25 °C and a volume of 20.0 mL. If the temperature is increased to 120 °C, and the piston expands to a volume of 80 mL, what is the new pressure inside the piston?

74. A gas inside a piston occupies a volume of 220 cm^3 at a pressure of 1.03 bars and a temperature of 100 °C. The gas is then cooled to room temperature (25.0 °C), and the volume of gas drops to 180 cm^3. What is the pressure of the gas under these new conditions?

75. A gas inside a balloon has a temperature of 293 K and a volume of 2.40 liters. The gas is cooled to 273 K, and pressure decreases from 790 mm Hg to 750 mm Hg. What is the new volume of the balloon? Report your answer to three significant digits.

76. In air-conditioning systems, compressed gases are allowed to expand, and this expansion results in cooling. A gas with a volume of 5.00 mL at a pressure of 8.0 bars at a temperature of 40 °C is allowed to expand to a volume of 15.0 mL at a pressure of 2.0 bars. What is the temperature of the gas after it expands?

77. What is the volume of two moles of gas, calculated at standard temperature and pressure?

78. What is the volume of 3.5 moles of gas, calculated at standard temperature and pressure?

79. At STP, how many moles of gas occupy a room with a volume of 5.0 m^3? (1 m^3 = 1,000 L)

80. At STP, how many moles of gas occupy a tank with a volume of 1.2 m^3? (1 m^3 = 1,000 L)

81. What is the pressure of 2.31 moles of gas at a temperature of 400.0 K and a volume of 3.5 liters?

82. What is the pressure of 12.5 moles of gas at a temperature of 360.0 K and a volume of 5.02 liters?

83. At what temperature does 4.0 moles of gas under a pressure of 1.0 atmosphere occupy a volume of 120.3 liters?

84. At what temperature does 1.3 moles of gas under a pressure of 1.2 atmospheres occupy a volume of 108.4 L?

85. A gas cylinder contains 80.0 grams of helium gas, occupying a volume of 20.0 liters at a temperature of 280 K. What is the pressure inside the cylinder?

86. What is the pressure of 14.0 moles of argon gas at a temperature of 0 °C and a volume of 800 mL?

87. What volume is occupied by 2 moles of nitrogen gas at a temperature of 73 °F and a pressure of 238.0 psi?

88. What volume would be required to store 150 moles of argon gas at 70 °F with a maximum pressure of 350.0 psi?

89. A gas cylinder contains 80.0 grams of carbon dioxide gas, occupying a volume of 20.0 liters at a temperature of 280 K. What is the pressure inside the cylinder?

90. How many grams of helium gas can a 12.0-liter cylinder at a temperature of 30 °C contain before the pressure exceeds 180.0 psi?

91. A 1.0-L cylinder of carbon dioxide has a pressure of 8.0 atmospheres and a temperature of 373 K. How many CO_2 molecules are in this cylinder?

92. What is the pressure caused by 1.5×10^{21} helium atoms occupying a volume of 0.0023 L at a temperature of 398 K?

93. A gas cylinder contains an unknown gas. It is found that a 1.90-gram sample of the gas occupies a volume of 2.30 liters at a pressure of 1.0 atmosphere and a temperature of 298 K.

a. How many moles are in the 1.90-gram sample?
b. What is the formula mass of this gas, in grams per mole?
c. Based on the formula mass, identify the gas as one of the following: helium, neon, argon, nitrogen, oxygen, or carbon dioxide.

94. A gas cylinder at a temperature of 298 K contains an unknown gas. When 48.0 grams of the gas are released from the cylinder, the pressure of the gas drops by 53.9 psi (3.67 atmospheres). How many moles of gas were released from the tank? What is the formula mass of this gas, in grams per mole? Identify the gas as one of the following: helium, neon, argon, nitrogen, oxygen, or carbon dioxide.

95. What does the term *partial pressure* mean?

96. A gas mixture contains 50% oxygen, 30% nitrogen, and 20% carbon dioxide by mole ratio. If the total pressure is 1.0 atmosphere, what is the partial pressure of oxygen, nitrogen, and carbon dioxide in the sample?

97. A gas mixture contains 40% methane and 60% oxygen by mole ratio. If the total pressure is 45 psi, what is the partial pressure of methane and oxygen in the sample?

98. A cylinder of gas contains 100.0 moles of nitrogen gas and 25.0 moles of oxygen gas. The overall pressure of the cylinder is 214.2 psi. Calculate the partial pressure of oxygen in the cylinder.

99. A 24.0-liter gas cylinder contains 8.0 moles of nitrogen gas and 2.0 moles of oxygen gas. If the temperature inside the cylinder is 0 °C, what is the total pressure inside the cylinder?

100. A balloon is filled with 10.9 grams of N_2, 9.6 grams of O_2, 2.2 grams of CO_2, and 0.04 grams of Ar. The pressure inside the balloon is 1.05 bars, and the temperature is 70 °F. What is the volume of the balloon?

101. An air sample contains 0.5% water vapor. If the total air pressure is 750 torr, what is the partial pressure due to the water vapor?

102. Imagine you are sitting in a room with a volume of 25,000 liters and a temperature of 72 °F. The total air pressure is 750 mm Hg, and the partial pressure from oxygen is 21% of the total pressure. How many grams of oxygen are in the room?

103. A mixture of nitrogen gas (N_2) and oxygen gas (O_2) has a pressure of 1.0 atmosphere at a temperature of 72 °F. At this temperature, which type of molecules is moving with more velocity?

104. The pressure that a gas exerts on a container is determined by the collisions the gas makes with the inside walls of the container. How would these collisions change (in frequency or magnitude) under the following conditions?

a. Keep the temperature constant and decrease the volume.
b. Decrease the temperature and decrease the volume.
c. Increase the temperature and decrease the volume.
d. Increase both the temperature and the volume.

10.5 Diffusion and Effusion

105. What is the difference between diffusion and effusion?

106. If you have a balloon filled with neon and a balloon filled with argon, which one will deflate the fastest? How do you know?

10.6 Gas Stoichiometry

107. The reaction of potassium metal with water produces hydrogen gas, as in the equation below. How many moles of hydrogen gas can be produced from the reaction of 5 moles of potassium? At STP, how many liters of hydrogen gas can be produced?

$$2\ K\ (s) + 2\ H_2O\ (aq) \rightarrow 2\ KOH\ (aq) + H_2\ (g)$$

108. Hydrogen and oxygen react explosively to form water, as in the reaction below. If a balloon containing 1.5 liters of hydrogen gas at 25 °C and a pressure of 1.0 atmosphere reacts with excess oxygen, how many grams of water can be produced?

$$2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (g)$$

109. Propane gas (C_3H_8) reacts with oxygen according to this balanced equation:

$$C_3H_8\ (g) + 5\ O_2\ (g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (g)$$

a. How many moles of water can be formed from 15.0 moles of propane?
b. How many moles of carbon dioxide can be formed from 15.0 moles of propane?
c. At STP, what volume of CO_2 would be produced from 15.0 moles of propane?
d. Are more moles of gas produced or consumed in this reaction?

110. The Born–Haber process is used to manufacture ammonia (NH_3) from nitrogen gas and hydrogen gas:

$$3\ H_2\ (g) + N_2\ (g) \rightarrow 2\ NH_3\ (g)$$

In this reaction,

a. How many moles of nitrogen are needed to react with 15 moles of hydrogen?
b. How many moles of ammonia can be produced from 15 moles of hydrogen?
c. At a temperature of 800 K and a pressure of 4.00 atm, how many liters of ammonia can be produced from 15.0 moles of hydrogen?
d. If this reaction proceeded to the right in a sealed container, would the pressure inside the container increase or decrease? How do you know?

111. Natural gas, CH_4, burns in oxygen as shown in the reaction below. A 1.0-liter cylinder containing CH_4 with a pressure of 2.1 atmospheres at a temperature of 298 K is completely used to power a portable stove.

$$CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (g)$$

a. How many moles of CH_4 were in the cylinder?
b. How many moles of oxygen gas were necessary to react with this amount of CH_4?
c. How many moles of water were produced in this reaction?
d. How many grams of carbon dioxide were produced in this reaction?

112. When a beverage is carbonated, the carbon dioxide reacts with water to produce carbonic acid, which gives soft drinks a bitterness that is balanced with a large amount of sugar. Over time, carbonic acid converts back into carbon dioxide and water:

$$H_2CO_3\ (aq) \rightarrow H_2O\ (l) + CO_2\ (g)$$

In this reaction, what volume of CO_2 can be produced from 2.0 grams of H_2CO_3 at standard temperature and pressure? (Standard temperature and pressure is 273 K and 1.0 atmosphere.)

113. Glucose fermentation takes place through the following reaction:

$$C_6H_{12}O_6\ (aq) \rightarrow 2\ C_2H_6O\ (aq) + 2\ CO_2\ (g)$$

If 2.05 grams of $C_6H_{12}O_6$ were placed in a sealed container having a temperature of 200 °C and a volume of 5.0 liters, and if this reaction went to completion, what would be the pressure from the carbon dioxide inside the container?

114. Ammonia, NH_3, is a colorless gas with a pungent odor. Its applications range from cleaning supplies to fertilizer, and it is a building block for the production of many industrial chemicals and consumer products. For example, the first step in preparing nitric acid involves the reaction of ammonia gas with oxygen gas to produce nitrogen monoxide and water, as shown in this reaction:

$$4\ NH_3\ (g) + 5\ O_2\ (g) \rightarrow 4\ NO\ (g) + 6\ H_2O\ (g)$$

a. If this reaction takes place in a sealed container, and the temperature inside the container is kept constant, will the pressure inside the container increase or decrease? How do you know?
b. If 170 grams of NH_3 react in this way, how many grams of H_2O can be produced?
c. If 170 grams of NH_3 react in this way, how many liters of H_2O can be produced, given a temperature of 373 K and a pressure of 1.00 atmosphere?

115. Octane, a component of gasoline, burns according to this equation:

$$2\ C_8H_{18}\ (l) + 25\ O_2\ (g) \rightarrow 16\ CO_2\ (g) + 18\ H_2O\ (g)$$

Octane has a formula mass of 114.26 g/mol and a density of 0.703 kg/L. At standard temperature and pressure, what volume of CO_2 is produced by the combustion of one liter of octane?

116. During photosynthesis, plants absorb sunlight and use this energy to convert carbon dioxide and water into simple sugars and oxygen. The process is exactly the reverse of combustion and is described by this equation:

$$6\ CO_2\ (g) + 6\ H_2O\ (l) \rightarrow C_6H_{12}O_6\ (s) + 6\ O_2\ (g)$$

At a temperature of 25 °C and a pressure of 1.00 atm, how many liters of carbon dioxide gas are consumed by the production of 1.00 kg of $C_6H_{12}O_6$?

Challenge Questions

117. Propane gas reacts with oxygen according to this balanced equation:

$$C_3H_8\ (g) + 5\ O_2\ (g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (g)$$

Assuming an air sample contains 21% oxygen (by volume), in which of the following mixtures will the propane and oxygen almost completely react, leaving almost no excess of either gas?

a. a mixture of 50% air and 50% propane
b. a mixture of 74% air and 26% propane
c. a mixture of 82% air and 18% propane
d. a mixture of 96% air and 4% propane

118. Octane, a component of gasoline, burns according to this equation:

$$2\ C_8H_{18}\ (l) + 25\ O_2\ (g) \rightarrow 16\ CO_2\ (g) + 18\ H_2O\ (g)$$

In an engine cycle, 0.030 g of C_8H_{18} is mixed with 0.30 L of air at a pressure of 3.2 atmospheres and a temperature of 450 K inside a piston. The spark plug detonates the gasoline, causing the explosive reaction to take place. What mass of CO_2 and what mass of H_2O can be produced in this engine cycle? Assume the air contains 21% oxygen. In what two ways does this reaction increase the pressure and push the piston?