Blood pressure is the force within arteries that drives the circulation of blood throughout the body. Blood pressure in the body is analogous to water pressure in a plumbing system. Just as water pressure pushes water through the pipes to faucets and fixtures throughout a house, blood pressure pushes blood to muscles and other tissues throughout the body. However, unlike the water pressure in a plumbing system—which is typically nearly constant—our blood pressure varies with each heartbeat. When the heart muscle contracts, blood pressure increases; between contractions it decreases. Systolic blood pressure is the peak pressure during a contraction, and diastolic blood pressure is the lowest pressure between contractions. Just as excessively high water pressure in a plumbing system can damage pipes, so too high blood pressure in a circulatory system can damage the heart and arteries, resulting in increased risk of stroke and heart attack.

Medical professionals usually measure blood pressure with an instrument called a sphygmomanometer—an inflatable cuff equipped with a pressure gauge—and a stethoscope. The cuff is wrapped around the patient’s arm and inflated with air. As air is pumped into the cuff, the pressure in the cuff increases. The cuff tightens around the arm and compresses the artery, momentarily stopping blood flow. The person measuring the blood pressure listens to the artery through the stethoscope while slowly releasing the pressure in the cuff. When the pressure in the cuff equals the systolic blood pressure (the peak pressure), a pulse is heard through the stethoscope. The pulse is the sound of blood getting through the compressed artery during a contraction of the heart. The pressure reading at that exact moment is the systolic blood pressure. As the pressure in the cuff continues to decrease, the blood can flow through the compressed artery even between contractions, so the pulsing sound stops. The pressure reading when the pulsing sound stops is the diastolic blood pressure (the lowest pressure).

A blood pressure measurement is usually reported as two pressures, in mmHg, separated by a slash. For example, a blood pressure measurement of 122/84 indicates that the systolic blood pressure is 122 mmHg and the diastolic blood pressure is 84 mmHg. Although the value of blood pressure can vary throughout the day, a healthy (or normal) value is usually considered to be below 120 mmHg for systolic and below 80 mmHg for diastolic (Table 5.2). High blood pressure, also called hypertension, entails the health risks mentioned previously.

Risk factors for hypertension include obesity, high salt (sodium) intake, high alcohol intake, lack of exercise, stress, a family history of high blood pressure, and age (blood pressure tends to increase as we get older). Mild hypertension can be managed with diet and exercise. Moderate to severe cases require doctor-prescribed medication.

### TABLE 5.2 Blood Pressure Ranges

<table>
<thead>
<tr>
<th>Blood Pressure</th>
<th>Systolic (mmHg)</th>
<th>Diastolic (mmHg)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hypotension</td>
<td>&lt;100</td>
<td>&lt;60</td>
</tr>
<tr>
<td>Normal</td>
<td>100–119</td>
<td>60–79</td>
</tr>
<tr>
<td>Prehypertension</td>
<td>120–139</td>
<td>80–89</td>
</tr>
<tr>
<td>Hypertension Stage 1</td>
<td>140–159</td>
<td>90–99</td>
</tr>
<tr>
<td>Hypertension Stage 2</td>
<td>&gt;160</td>
<td>&gt;100</td>
</tr>
</tbody>
</table>

5.3 The Simple Gas Laws: Boyle’s Law, Charles’s Law, and Avogadro’s Law

We have learned about pressure and its characteristics. We now broaden our discussion to include the four basic properties of a gas sample: pressure (P), volume (V), temperature (T), and amount in moles (n). These properties are interrelated—when one changes, it affects the others. The simple gas laws describe the relationships between pairs of these properties. For example, how does volume vary with pressure at constant temperature and amount of gas, or with temperature at constant pressure and amount of gas? We can
elucidate these relationships by conducting experiments in which two of the four basic properties are held constant in order to determine the relationship between the other two. We can then express the results of the experiments as laws, called the simple gas laws.

**Boyle’s Law: Volume and Pressure**

In the early 1660s, the pioneering English scientist Robert Boyle (1627–1691) and his assistant Robert Hooke (1635–1703) used a J-tube (Figure 5.6) to measure the volume of a sample of gas at different pressures. They trapped a sample of air in the J-tube and added mercury to increase the pressure on the gas. Boyle and Hooke observed an inverse relationship between volume and pressure—an increase in one results in a decrease in the other—as shown in Figure 5.7. This relationship is now known as **Boyle’s law**.

Boyle’s law:

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

(constant \( T \) and \( n \))

Boyle’s law follows from the idea that pressure results from the collisions of the gas particles with the walls of their container. If the volume of a gas sample is decreased, the same number of gas particles is crowded into a smaller volume, resulting in more collisions with the walls and therefore an increase in the pressure (Figure 5.8).

Scuba divers learn about Boyle’s law during certification because it explains why they should not ascend toward the surface without continuous breathing. For every 10 m of depth that a diver descends in water, she experiences an additional 1 atm of pressure due to the weight of the water above her (Figure 5.9). The pressure regulator used in scuba diving delivers air into the diver’s lungs at a pressure that matches the external pressure; otherwise the diver could not inhale the air (see *Chemistry in Your Day: Extra-long Snorkels* on page 187). For example, when a diver is 20 m below the surface, the regulator delivers air at a pressure of 3 atm to match the 3 atm of pressure around the diver (1 atm due to normal atmospheric pressure and 2 additional atmospheres due to the weight of the water at 20 m). Suppose that a diver inhaled a lungful of air at a pressure of 3 atm and swam quickly to the surface (where the pressure is 1 atm) while holding her breath. What would happen to the volume of air in her lungs? Since the pressure decreases by a factor of 3, the volume of the air in her lungs would increase by a factor of 3—a dangerous situation. Of course, the volume increase in the diver’s lungs would be so great that she would not be able to hold her breath all the way to the surface—the air would...
5.3 The Simple Gas Laws: Boyle's Law, Charles's Law, and Avogadro's Law

**Volume versus Pressure: A Molecular View**

![Diagram showing Boyle's Law]

force itself out of her mouth but probably not before the expanded air severely damaged her lungs, possibly killing her. Consequently, the most important rule in diving is *never hold your breath*. To avoid such catastrophic results, divers must ascend slowly and breathe continuously, allowing the regulator to bring the air pressure in their lungs back to 1 atm by the time they reach the surface.

We can use Boyle's law to calculate the volume of a gas following a pressure change or the pressure of a gas following a volume change as long as the temperature and the amount of gas remain constant. For these types of calculations, we write Boyle's law in a slightly different way.

If we multiply both sides by $P$, we get

$$PV = \text{constant}$$
This relationship indicates that if the pressure increases, the volume decreases, but the product $P \times V$ always equals the same constant. For two different sets of conditions, we can say that

$$P_1 V_1 = constant = P_2 V_2$$

or

$$P_1 V_1 = P_2 V_2$$  \[5.2\]

where $P_1$ and $V_1$ are the initial pressure and volume of the gas and $P_2$ and $V_2$ are the final volume and pressure.

**EXAMPLE 5.2 Boyle’s Law**

A cylinder equipped with a movable piston has a volume of 7.25 L under an applied pressure of 4.52 atm. What is the volume of the cylinder if we decrease the applied pressure to 1.21 atm?

To solve the problem, first solve Boyle’s law (Equation 5.2) for $V_2$ and then substitute the given quantities to calculate $V_2$.

**SOLUTION**

$$P_1 V_1 = P_2 V_2$$

$$V_2 = \frac{P_1}{P_2} V_1$$

$$= \frac{4.52 \text{ atm}}{1.21 \text{ atm}} \times 7.25 \text{ L}$$

$$= 27.1 \text{ L}$$

**FOR PRACTICE 5.2**

A snorkeler takes a syringe filled with 16 mL of air from the surface, where the pressure is 1.0 atm, to an unknown depth. The volume of the air in the syringe at this depth is 7.5 mL. What is the pressure at this depth? If the pressure increases by 1 atm for every additional 10 m of depth, how deep is the snorkeler?

**Charles’s Law: Volume and Temperature**

Suppose we keep the pressure of a gas sample constant and measure its volume at a number of different temperatures. Figure 5.10 shows the results of several such measurements. From the plot we can see a relationship between volume and temperature: the volume of...
Several episodes of The Flintstones cartoon featured Fred Flintstone and Barney Rubble snorkeling. Their snorkels, however, were not the modern kind, but long reeds that stretched from the surface of the water down to many meters of depth. Fred and Barney swam around in deep water while breathing air provided to them by these extra-long snorkels. Would this work? Why do people bother with scuba diving equipment if they could instead simply use 10-m snorkels as Fred and Barney did?

As we saw in Section 5.1, when we breathe, we expand the volume of our chest cavity, reducing the pressure on the outer surface of the lungs to less than 1 atm (Boyle's law). Because of this pressure differential, the lungs expand, the pressure in them falls, and air from outside of our lungs then flows into them. Extra-long snorkels do not work because of the pressure exerted by water at depth. A diver at 10 m experiences an external pressure of 2 atm. This is more than the muscles of the chest cavity can overcome—the chest cavity and lungs are compressed, resulting in an air pressure within them of more than 1 atm. If the diver had a snorkel that went to the surface—where the air pressure is 1 atm—air would flow out of his lungs (from greater pressure to less pressure), not into them. It would be impossible to breathe.

**Question**

A diver takes a balloon with a volume of 2.5 L from the surface, where the pressure is 1.0 atm, to a depth of 20 m, where the pressure is 3.0 atm. What happens to the volume of the balloon? What if the end of the submerged balloon were on a long pipe that went to the surface and was attached to another balloon? Which way would air flow as the diver descended?

A gas increases with increasing temperature. Looking at the plot more closely, however, reveals more—volume and temperature are linearly related. If two variables are linearly related, then plotting one against the other produces a straight line.

Another interesting feature emerges if we extend or extrapolate the line in the plot backwards from the lowest measured temperature. The dotted extrapolated line shows that the gas should have a zero volume at -273.15 °C. Recall from Chapter 1 that -273.15 °C corresponds to 0 K (zero on the Kelvin scale), the coldest possible temperature. The extrapolated line shows that below -273.15 °C, the gas would have a negative volume, which is physically impossible. For this reason, we refer to 0 K as *absolute zero*—colder temperatures do not exist.

The first person to carefully quantify the relationship between the volume of a gas and its temperature was J. A. C. Charles (1746–1823), a French mathematician and physicist. Charles was interested in gases and was among the first people to ascend in a hydrogen-filled balloon. The direct proportionality between volume and temperature is named *Charles's law* after him.

Charles's law: \[ V \propto T \] (constant \( P \) and \( n \))

When the temperature of a gas sample is increased, the gas particles move faster; collisions with the walls are more frequent, and the force exerted with each collision is greater. The only way for the pressure (the force per unit area) to remain constant is for
A hot-air balloon floats because the hot air is less dense than the surrounding cold air.

If we place a balloon into liquid nitrogen (77 K), it shrivels up as the air within it cools and occupies less volume at the same external pressure.

The gas to occupy a larger volume, so that collisions become less frequent and occur over a larger area (Figure 5.11a).

Charles’s law explains why the second floor of a house is usually warmer than the ground floor. According to Charles’s law, when air is heated, its volume increases, resulting in a lower density. The warm, less dense air tends to rise in a room filled with colder, denser air. Similarly, Charles’s law explains why a hot-air balloon can take flight. The gas that fills a hot-air balloon is warmed with a burner, increasing its volume and lowering its density, and causing it to float in the colder, denser surrounding air.

You can experience Charles’s law directly by holding a partially inflated balloon over a warm toaster. As the air in the balloon warms, you can feel the balloon expanding. Alternatively, you can put an inflated balloon into liquid nitrogen and see that it becomes smaller as it cools.

We can use Charles’s law to calculate the volume of a gas following a temperature change or the temperature of a gas following a volume change as long as the pressure and the amount of gas are constant. For these calculations, we rearrange Charles’s law as follows:

\[
\text{Since } V \propto T, \text{ then } V = \text{constant } \times T
\]

If we divide both sides by \( T \), we get

\[
\frac{V}{T} = \text{constant}
\]

If the temperature increases, the volume increases in direct proportion so that the quotient, \( V/T \), is always equal to the same constant. So, for two different measurements, we can say that

\[
\frac{V_1}{T_1} = \text{constant} = \frac{V_2}{T_2},
\]

or

\[
\frac{V_1}{T_1} = \frac{V_2}{T_2}
\]

where \( V_1 \) and \( T_1 \) are the initial volume and temperature of the gas and \( V_2 \) and \( T_2 \) are the final volume and temperature. \textit{The temperatures must always be expressed in kelvins (K)}, because, as you can see in Figure 5.10, the volume of a gas is directly proportional to its absolute temperature, not its temperature in °C. For example, doubling the temperature of a gas sample from 1 °C to 2 °C does not double its volume, but doubling the temperature from 200 K to 400 K does.
EXAMPLE 5.3 Charles’s Law

A sample of gas has a volume of 2.80 L at an unknown temperature. When the sample is submerged in ice water at $T = 0.00 \, ^\circ C$, its volume decreases to 2.57 L. What was its initial temperature (in K and in °C)?

To solve the problem, first solve Charles’s law for $T_1$.

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

Before you substitute the numerical values to calculate $T_1$, you must convert the temperature to kelvins (K). Remember, gas law problems must always be worked with Kelvin temperatures.

Substitute $T_2$ and the other given quantities to calculate $T_1$.

$$T_1 \frac{V_1}{V_2} = \frac{2.80 \, L}{2.57 \, L} = 273.15 \, K$$

$$T_1 = \frac{2.80 \, L}{2.57 \, L} \cdot 273.15 \, K = 297.6 \, K$$

Calculate $T_1$ in °C by subtracting 273 from the value in kelvins.

$$T_1 (\, ^\circ C) = 297.6 - 273.15 = 24 \, ^\circ C$$

FOR PRACTICE 5.3

A gas in a cylinder with a moveable piston has an initial volume of 88.2 mL. If we heat the gas from 35 °C to 155 °C, what is its final volume (in mL)?

Avogadro’s Law: Volume and Amount (in Moles)

So far, we have learned the relationships between volume and pressure, and volume and temperature, but we have considered only a constant amount of a gas. What happens when the amount of gas changes? The volume of a gas sample (at constant temperature and pressure) as a function of the amount of gas (in moles) in the sample is shown in Figure 5.12. We can see that the relationship between volume and amount is linear. As we might expect, extrapolation to zero moles shows zero volume. This relationship, first stated formally by Amadeo Avogadro, is called Avogadro’s law:

Avogadro’s law: $V \propto n$ (constant $T$ and $P$)

When the amount of gas in a sample increases at constant temperature and pressure, its volume increases in direct proportion because the greater number of gas particles fill more space.

You experience Avogadro’s law when you inflate a balloon. With each exhaled breath, you add more gas particles to the inside of the balloon, increasing its volume. We can use Avogadro’s law to calculate the volume of a gas following a change in the amount of the gas as long as the pressure and temperature of the gas are constant. For these types of calculations, we express Avogadro’s law as

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \quad \quad \quad [5.4]$$

where $V_1$ and $n_1$ are the initial volume and number of moles of the gas and $V_2$ and $n_2$ are the final volume and number of moles. In calculations, we use Avogadro’s law in a manner similar to the other gas laws, as demonstrated in the following example.
EXAMPLE 5.4 Avogadro’s Law

A 4.65-L sample of helium gas contains 0.225 mol of helium. How many additional moles of helium gas must we add to the sample to obtain a volume of 6.48 L? Assume constant temperature and pressure.

To solve the problem, first solve Avogadro’s law for \( n_2 \). Then substitute the given quantities to calculate \( n_2 \).

Since the balloon already contains 0.225 mol of gas, calculate the amount of gas to add by subtracting 0.225 mol from the value you calculated for \( n_2 \). (In Chapter 1, we introduced the practice of underlining the least (rightmost) significant digit of intermediate answers, but not rounding the final answer until the very end of the calculation. We continue that practice in this chapter. However, in order to avoid unnecessary notation, we will not carry additional digits in cases, such as this one, where doing so would not affect the final answer.)

SOLUTION

\[
\frac{V_1}{n_1} = \frac{V_2}{n_2}
\]

\[
n_2 = \frac{V_2}{V_1} n_1
\]

\[
= \frac{6.48 \text{ L}}{4.65 \text{ L}} \times 0.225 \text{ mol}
\]

\[
= 0.314 \text{ mol}
\]

moles to add = 0.314 mol - 0.225 mol

\[
= 0.089 \text{ mol}
\]

FOR PRACTICE 5.4

A chemical reaction occurring in a cylinder equipped with a moveable piston produces 0.621 mol of a gaseous product. If the cylinder contained 0.120 mol of gas before the reaction and had an initial volume of 2.18 L, what was its volume after the reaction? (Assume constant pressure and temperature and that the initial amount of gas completely reacts.)

5.4 The Ideal Gas Law

The relationships that we have learned so far can be combined into a single law that encompasses all of them. So far, we know that

\[
V \propto \frac{1}{P} \quad \text{(Boyle’s law)}
\]

\[
V \propto T \quad \text{(Charles’s law)}
\]

\[
V \propto n \quad \text{(Avogadro’s law)}
\]

Combining these three expressions, we get

\[
V \propto \frac{nT}{P}
\]

The volume of a gas is directly proportional to the number of moles of gas and to the temperature of the gas, but is inversely proportional to the pressure of the gas. We can replace the proportionality sign with an equals sign by incorporating \( R \), a proportionality constant called the *ideal gas constant*:

\[
V = \frac{RnT}{P}
\]

Rearranging, we get

\[
P V = nRT \quad \text{[5.5]}
\]
This equation is the ideal gas law, and a hypothetical gas that exactly follows this law is an ideal gas. The value of \( R \), the ideal gas constant, is the same for all gases and has the following value:

\[
R = \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}
\]

The ideal gas law contains within it the simple gas laws that we have learned. For example, recall that Boyle’s law states that \( V \propto 1/P \) when the amount of gas (\( n \)) and the temperature of the gas (\( T \)) are kept constant. We can rearrange the ideal gas law as follows:

\[
PV = nRT
\]

First, divide both sides by \( P \):

\[
V = \frac{nRT}{P}
\]

Then put the variables that are constant, along with \( R \), in parentheses:

\[
V = (nRT) \frac{1}{P}
\]

Since \( n \) and \( T \) are constant in this case, and since \( R \) is always a constant, we can write

\[
V \propto (\text{constant}) \times \frac{1}{P}
\]

which means that \( V \propto 1/P \).

The ideal gas law also shows how other pairs of variables are related. For example, from Charles’s law we know that \( V \propto T \) at constant pressure and constant number of moles. But what if we heat a sample of gas at constant volume and constant number of moles? This question applies to the warning labels on aerosol cans such as hair spray or deodorants. These labels warn against excessive heating or incineration of the can, even after the contents are used up. Why? An “empty” aerosol can is not really empty but contains a fixed amount of gas trapped in a fixed volume. What would happen if you were to heat the can? Let’s rearrange the ideal gas law to clearly see the relationship between pressure and temperature at constant volume and constant number of moles:

\[
PV = nRT
\]

\[
P = \frac{nRT}{V} = \left( \frac{nR}{V} \right)T
\]

Since \( n \) and \( V \) are constant and since \( R \) is always a constant:

\[
P = (\text{constant}) \times T
\]

This relationship between pressure and temperature is also known as Gay-Lussac’s law. As the temperature of a fixed amount of gas in a fixed volume increases, the pressure increases. In an aerosol can, this pressure increase can blow the can apart, which is why aerosol cans should not be heated or incinerated. They might explode.

The ideal gas law can also be used to determine the value of any one of the four variables (\( P, V, n, \) or \( T \)) given the other three. To do so, each of the quantities in the ideal gas law must be expressed in the units within \( R \):

- pressure (\( P \)) in atm
- volume (\( V \)) in L
- moles (\( n \)) in mol
- temperature (\( T \)) in K

\[\text{The ideal gas law contains the simple gas laws within it.}\]

\[\text{The labels on most aerosol cans warn against incineration. Since the volume of the can is constant, an increase in temperature causes an increase in pressure and possibly an explosion.}\]
EXAMPLE 5.5  Ideal Gas Law I

Calculate the volume occupied by 0.845 mol of nitrogen gas at a pressure of 1.37 atm and a temperature of 315 K.

SORT  The problem gives you the number of moles of nitrogen gas, the pressure, and the temperature. You are asked to find the volume.

GIVEN:  \( n = 0.845 \text{ mol} \),  \( P = 1.37 \text{ atm} \),  \( T = 315 \text{ K} \)

FIND:  \( V \)

STRATEGIZE  You are given three of the four variables (\( P \), \( T \), and \( n \)) in the ideal gas law and asked to find the fourth (\( V \)). The conceptual plan shows how the ideal gas law provides the relationship between the known quantities and the unknown quantity.

CONCEPTUAL PLAN

\[ PV = nRT \]

RELATIONSHIP USED  \( PV = nRT \) (ideal gas law)

SOLVE To solve the problem, first solve the ideal gas law for \( V \).

\[
V = \frac{nRT}{P}
\]

Then substitute the given quantities to calculate \( V \).

\[
V = \frac{0.845 \text{ mol} \times 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 315 \text{ K}}{1.37 \text{ atm}} = 15.9 \text{ L}
\]

CHECK The units of the answer are correct. The magnitude of the answer (15.9 L) makes sense because, as you will see in the next section, one mole of an ideal gas under standard conditions (273 K and 1 atm) occupies 22.4 L. Although these are not standard conditions, they are close enough for a ballpark check of the answer. Since this gas sample contains 0.845 mol, a volume of 15.9 L is reasonable.

FOR PRACTICE 5.5

An 8.50-L tire contains 0.552 mol of gas at a temperature of 305 K. What is the pressure (in atm and psi) of the gas in the tire?

EXAMPLE 5.6  Ideal Gas Law II

Calculate the number of moles of gas in a 3.24-L basketball inflated to a total pressure of 24.3 psi at 25 °C. (Note: The total pressure is not the same as the pressure read on a pressure gauge such as the kind used for checking a car or bicycle tire. That pressure, called the gauge pressure, is the difference between the total pressure and atmospheric pressure. In this case, if atmospheric pressure is 14.7 psi, the gauge pressure would be 9.6 psi. However, for calculations involving the ideal gas law, you must use the total pressure of 24.3 psi.)

SORT  The problem gives you the pressure, the volume, and the temperature. You are asked to find the number of moles of gas.

GIVEN:  \( P = 24.3 \text{ psi} \),  \( V = 3.24 \text{ L} \),  \( (\text{°C}) = 25 \text{ °C} \)

FIND:  \( n \)

STRATEGIZE  The conceptual plan shows how the ideal gas law provides the relationship between the given quantities and the quantity to be found.

CONCEPTUAL PLAN

\[ PV = nRT \]

RELATIONSHIP USED  \( PV = nRT \) (ideal gas law)
5.5 Applications of the Ideal Gas Law: Molar Volume, Density, and Molar Mass of a Gas

SOLVE To solve the problem, first solve the ideal gas law for \( n \).

Before substituting into the equation, convert \( P \) and \( T \) into the correct units.

Finally, substitute into the equation and calculate \( n \).

SOLUTION

\[
P V = n R T
\]

\[
n = \frac{P V}{R T}
\]

\[
P = 24.3 \text{ psi} \times \frac{1 \text{ atm}}{14.7 \text{ psi}} = 1.6531 \text{ atm}
\]

(Since rounding the intermediate answer would result in a slightly different final answer, we mark the least significant digit in the intermediate answer, but don’t round until the end.)

\[
T \text{ (K) } = 25 + 273 = 298 \text{ K}
\]

\[
n = \frac{1.6531 \text{ atm} \times 3.24 \text{ L}}{0.08206 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \times 298 \text{ K}} = 0.219 \text{ mol}
\]

CHECK The units of the answer are correct. The magnitude of the answer (0.219 mol) makes sense because, as you will see in the next section, one mole of an ideal gas under standard conditions (273 K and 1 atm) occupies 22.4 L. At a pressure that is 65% higher than standard conditions, the volume of 1 mol of gas would be proportionally lower. Since this gas sample occupies 3.24 L, the answer of 0.219 mol is reasonable.

FOR PRACTICE 5.6

What volume does 0.556 mol of gas occupy at a pressure of 715 mmHg and a temperature of 58 °C?

FOR MORE PRACTICE 5.6

Find the pressure in mmHg of a 0.133-g sample of helium gas in a 648-mL container at a temperature of 32 °C.

5.5 Applications of the Ideal Gas Law: Molar Volume, Density, and Molar Mass of a Gas

We just examined how we can use the ideal gas law to calculate one of the variables (\( P, V, T, \) or \( n \)) given the other three. We now turn to three other applications of the ideal gas law: molar volume, density, and molar mass.

Molar Volume at Standard Temperature and Pressure

The volume occupied by one mole of a substance is its molar volume. For gases, we often specify the molar volume under conditions known as standard temperature \((T = 0 \text{ °C or 273 K})\) and pressure \((P = 1.00 \text{ atm})\), abbreviated as STP. Using the ideal gas law, we can determine that the molar volume of ideal gas at STP is

\[
V = \frac{n R T}{P} = \frac{1.00 \text{ mol} \times 0.08206 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \times 273 \text{ K}}{1.00 \text{ atm}} = 22.4 \text{ L}
\]

The molar volume of 22.4 L only applies at STP.

\[\text{△ One mole of any gas occupies approximately 22.4 L at standard temperature (273 K) and pressure (1.0 atm).}\]