# CHAPTER 11
## GASES

It's Monday morning, and Lilia is walking out of the chemistry building, thinking about the introductory lecture on gases that her instructor just presented. Dr. Scanlon challenged the class to try to visualize gases in terms of the model she described, so Lilia looks at her hand and tries to picture the particles in the air bombarding her skin at a rate of $10^{23}$ collisions per second. Lilia has high hopes that a week of studying gases will provide her with answers to the questions her older brothers and sisters posed to her the night before at a family dinner.

When Ted, who is a mechanic for a Formula One racing team, learned that Lilia was going to be studying gases in her chemistry class, he asked her to find out how to calculate gas density. He knows that when the density of the air changes, he needs to adjust the car's brakes and other components to improve its safety and performance. John, who is an environmental scientist, wanted to be reminded why balloons that carry his scientific instruments into the upper atmosphere expand as they rise. Amelia is an artist who recently began to add neon lights to her work. After bending the tubes into the desired shape, she fills them with gas from a high pressure cylinder. She wanted to know how to determine the number of tubes she can fill with one cylinder.

Lilia's sister Rebecca, the oldest, is a chemical engineer who could answer Ted's, John's, and Amelia's questions, but to give Lilia an opportunity to use her new knowledge, she keeps quiet except to describe a gas-related issue of her own. Rebecca is helping to design an apparatus in which two gases will react at high temperature, and her responsibility is to equip the reaction vessel with a valve that will keep the pressure from rising to dangerous levels. She started to explain to Lilia why increased temperature leads to increased pressure, but when Lilia asked what gas pressure was and what caused it, Rebecca realized that she had better save her explanation for the next family dinner. Lilia (like yourself) will learn about gas pressures and many other gas related topics by reading Chapter 11 of her textbook carefully and listening closely in lecture.

The gas particles in the air around us are constantly colliding with our skin.

## Review Skills

The presentation of information in this chapter assumes that you can already perform the tasks listed below. You can test your readiness to proceed by answering the Review Questions at the end of the chapter. This might also be a good time to read the Chapter Objectives, which precede the Review Questions.

- Convert between temperatures in the Celsius and Kelvin scales. (Section 2.6)
- Describe the particle nature of gases. (Section 3.1)
- Convert the amount of one substance in a given reaction to the amount of another substance in the same reaction, whether the amounts are described by mass of pure substance or volume of a solution containing a given molarity of one of the substances. (Sections 10.1 and 10.3)
- Given an actual yield and a theoretical yield for a chemical reaction (or enough information to calculate a theoretical yield), calculate the percent yield for the reaction. (Section 10.2)

## 11.1 Gases and Their Properties

If you want to understand how gases behave—such as why fresh air rushes into your lungs when certain chest muscles contract or how gases in a car's engine move the pistons and power the car—you need a clear mental image of the model chemists use to explain the properties of gases and the relationships between them. The model was introduced in Section 3.1, but we'll be adding some new components to it in the review presented here.
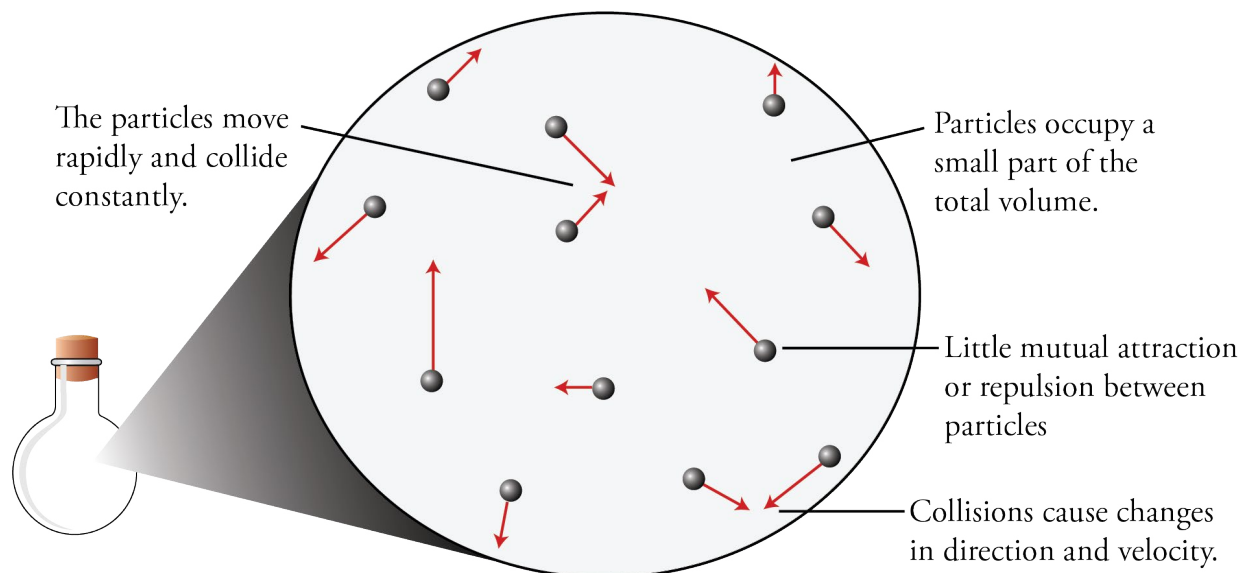
Gases consist of tiny particles widely spaced (Figure 11.1). Under typical conditions, the average distance between gas particles is about ten times their diameter. Because of these large distances, the volume occupied by the particles themselves is very small compared to the volume of the empty space around them. For a gas at room temperature and pressure, the gas particles themselves occupy about 0.1% of the total volume. The other 99.9% of the total volume is empty space (whereas in liquids and solids, about 70% of the volume is occupied by particles). Because of the large distances between gas particles, the attractions or repulsions among them are weak.

OBJECTIVE 2

The particles in a gas are in rapid and continuous motion. For example, the average velocity of nitrogen molecules, $N_2$, at 20 °C is about 500 m/s. As the temperature of a gas increases, the particles' velocity increases. The average velocity of nitrogen molecules at 100 °C is about 575 m/s.

The particles in a gas are constantly colliding with the walls of the container and with each other. Because of these collisions, the gas particles are constantly changing their direction of motion and their velocity. In a typical situation, a gas particle moves

OBJECTIVE 3

a very short distance between collisions. For example, oxygen, $O_2$, molecules at normal temperatures and pressures move an average of $10^{-7}$ m between collisions.

**Figure 11.1**
Particles of a Gas



The particles move rapidly and collide constantly.

Particles occupy a small part of the total volume.

Little mutual attraction or repulsion between particles

Collisions cause changes in direction and velocity.

## Ideal Gases

The model described above applies to *real* gases, but chemists often simplify the model further by imagining the behavior of an *ideal* gas. An ideal gas differs from a real gas in that

- The particles are assumed to be point masses, that is, particles that have a mass but occupy no volume.
- There are no attractive or repulsive forces at all between the particles.

When we add these assumptions to our model for gases, we call it the **ideal gas model**. As the name implies, the ideal gas model describes an "ideal" of gas behavior that is only approximated by reality. Nevertheless, the model succeeds in explaining and predicting the behavior of typical gases under typical conditions. In fact, some actual gases do behave very much in accordance with the model, and scientists may call them **ideal gases**. The ideal gas assumptions make it easier for chemists to describe the relationships between the properties of gases and allow us to calculate values for these properties.

## Properties of Gases

The ideal gas model is used to predict changes in four related gas properties: volume, number of particles, temperature, and pressure. Volumes of gases are usually described in liters, L, or cubic meters, $m^3$, and numbers of particles are usually described in moles, mol. Although gas temperatures are often measured with thermometers that report temperatures in degrees Celsius, °C, scientists generally use Kelvin temperatures for calculations. Remember that you can convert between degrees Celsius, °C, and kelvins, K, using the following equations.

? K = °C + 273.15

? °C = K − 273.15

To understand gas pressure, picture a typical gas in a closed container. Each time a gas particle collides with and ricochets off one of the walls of its container, it exerts a force against the wall. The sum of the forces of these ongoing collisions of gas particles against all the container's interior walls creates a continuous pressure upon those walls. **Pressure** is force divided by area.

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

$$\text{Gas pressure} = \frac{\text{Force due to particle collisions with the walls}}{\text{Area of the walls}}$$

The accepted SI unit for gas pressure is the pascal, Pa. A pascal is a very small amount of pressure, so the kilopascal, kPa, is more commonly used. Other units used to describe gas pressure are the atmosphere (atm), torr, and millimeter of mercury (mmHg). The relationships between these pressure units are

1 atm = 101,325 Pa = 101.325 kPa = 760 mmHg = 760 torr

1 bar = 100 kPa = 0.9869 atm = 750.1 mmHg

The numbers in these relationships come from definitions, so they are all exact. At sea level on a typical day, the atmospheric pressure is about 101 kPa, or about 1 atm.

In calculations, the variables *P*, *T*, *V*, and *n* are commonly used to represent pressure, temperature, volume, and moles of gas.

## Discovering the Relationships Between Properties

If we want to explain why a weather balloon carrying instruments into the upper atmosphere expands as it rises, we need to consider changes in the properties of the gases (pressure, volume, temperature, or number of gas particles) inside and outside the balloon. For example, as the balloon rises, the pressure outside of it, called the atmospheric pressure, decreases. But, there are also variations in temperature, and the balloon might have small leaks that change the number of gas particles it contains.

In a real situation, pressure, temperature, and number of gas particles may all be changing, and predicting the effect of such a blend of changing properties on gas volume is tricky. Therefore, before we tackle predictions for real world situations, such as the weather balloon, we will consider simpler systems in which two of the four gas properties are held constant, a third property is varied, and the effect of this variation on the fourth property is observed. For example, it is easier to understand the relationship between volume and pressure if the number of gas particles and temperature are held constant. The volume can be varied, and the effect this has on the pressure can be measured. An understanding of the relationships between gas properties in controlled situations will help us to explain and predict the effects of changing gas properties in more complicated, real situations.

Figure 11.2 shows a laboratory apparatus that can be used to demonstrate all the relationships we are going to be discussing. It consists of a cylinder with a movable piston, a thermometer, a pressure gauge, and a valve through which gas may be added to the cylinder's chamber or removed from it.

**Figure 11.2**
Apparatus Used to Demonstrate Relationships
Between the Properties of Gases
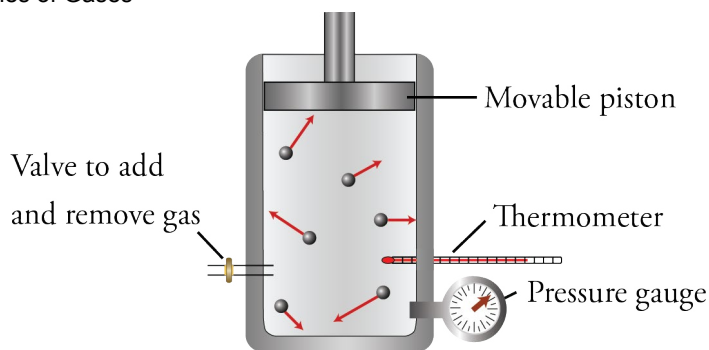


## The Relationship Between Volume and Pressure

OBJECTIVE 10A

Figure 11.3 shows how our demonstration apparatus would be used to determine the relationship between gas volume and pressure. While holding the number of gas particles constant (by closing the valve) and holding the temperature constant (by allowing heat to transfer in or out so that the apparatus remains the same temperature as the surrounding environment), we move the piston to change the volume, and then we observe the change in pressure. When we decrease the gas volume, the pressure gauge on our system shows us that the gas pressure increases. When we increase the gas volume, the gauge shows that the pressure goes down.

OBJECTIVE 10A

Decreased volume → Increased pressure
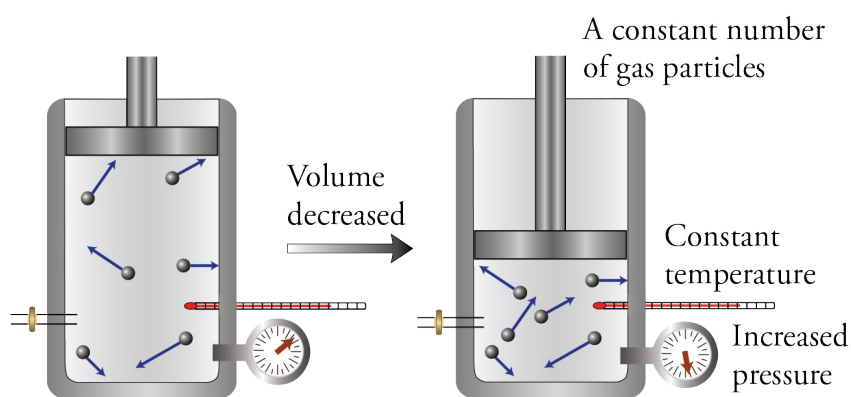
Increased volume → Decreased pressure

For an ideal gas (in which the particles occupy no volume and experience no attractions or repulsions), gas pressure and volume are inversely proportional. This means that if the temperature and the number of gas particles are constant and if the volume is decreased to one-half its original value, the pressure of the gas will double. If the volume is doubled, the pressure decreases to one-half its original value. The following expression summarizes this inverse relationship:

$$P \propto \frac{1}{V} \quad \text{if n and T are constant}$$

Real gases deviate somewhat from this mathematical relationship, but the general trend of increased pressure with decreased volume (or decreased pressure with increased volume) is true for any gas.

   The observation that *the pressure of an ideal gas is inversely proportional to the volume it occupies if the number of gas particles and the temperature are constant* is a statement of **Boyle's Law**. This relationship can be explained in the following way. When the volume of the chamber decreases but the number of gas particles remains constant, there is an increase in the concentration (number of particles per liter) of the gas. This leads to an increase in the number of particles near any given area of the container walls at any time and to an increase in the number of collisions against the walls per unit area in a given time. More collisions mean an increase in the force per unit area, or pressure, of the gas. The logic sequence presented in Figure 11.3 summarizes this explanation. The arrows in the logic sequence can be read as "leads to." Take the time to read the sequence carefully to confirm that each phrase leads logically to the next.
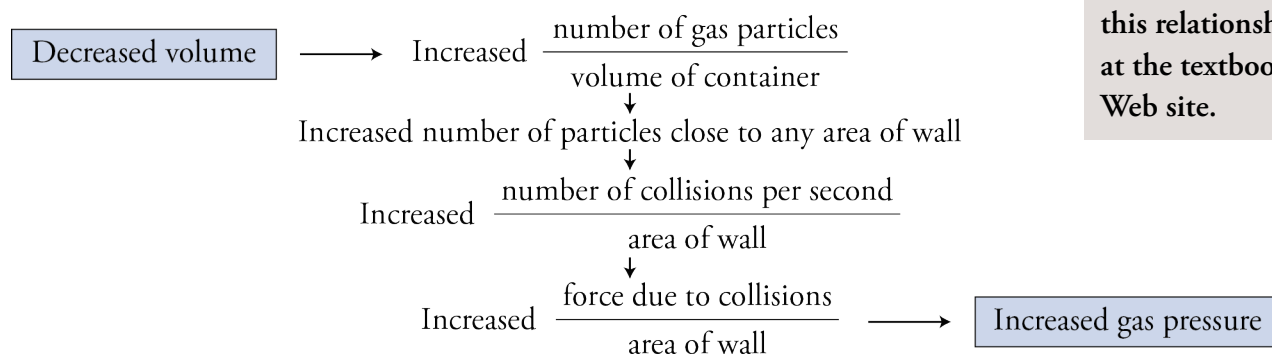


A constant number of gas particles

Volume decreased

Constant temperature

Increased pressure

**Figure 11.3
Relationship Between Volume and Pressure**
Decreased volume leads to increased pressure if the number of gas particles and the temperature are constant.

OBJECTIVE 10A

You can see an animation that demonstrates this relationship at the textbook's Web site.

| Decreased volume | $\longrightarrow$ | Increased $\dfrac{\text{number of gas particles}}{\text{volume of container}}$ |

$\downarrow$

Increased number of particles close to any area of wall

$\downarrow$

Increased $\dfrac{\text{number of collisions per second}}{\text{area of wall}}$

$\downarrow$

Increased $\dfrac{\text{force due to collisions}}{\text{area of wall}}$ $\longrightarrow$ Increased gas pressure

## The Relationship Between Pressure and Temperature

OBJECTIVE 10B

In order to examine the relationship between pressure and temperature, we must adjust our demonstration apparatus so that the other two properties (number of gas particles and volume) are held constant. This can be done by locking the piston so it cannot move and closing the valve tightly so that no gas leaks in or out (Figure 11.4). When the temperature of a gas trapped inside the chamber is increased, the measured pressure increases. When the temperature is decreased, the pressure decreases.

OBJECTIVE 10B

Increased temperature  →  Increased pressure
Decreased temperature  →  Decreased pressure

We can explain the relationship between temperature and pressure using our model for gas. Increased temperature means increased motion of the particles. If the particles are moving faster in the container, they will collide with the walls more often and with greater force per collision. This leads to a greater overall force pushing on the walls and to a greater force per unit area or pressure (Figure 11.4).

If the gas is behaving like an ideal gas, a doubling of the Kelvin temperature doubles the pressure. If the temperature decreases to 50% of the original Kelvin temperature, the pressure decreases to 50% of the original pressure. This relationship can be expressed by saying that *the pressure of an ideal gas is directly proportional to the Kelvin temperature of the gas if the volume and the number of gas particles are constant*. This relationship is sometimes called **Gay-Lussac's Law**.

OBJECTIVE 10B

$P \propto T$   if n and V are constant

**Figure 11.4**
**Relationship Between Temperature and Pressure**
Increased temperature leads to increased pressure if the number of gas particles and volume are constant.
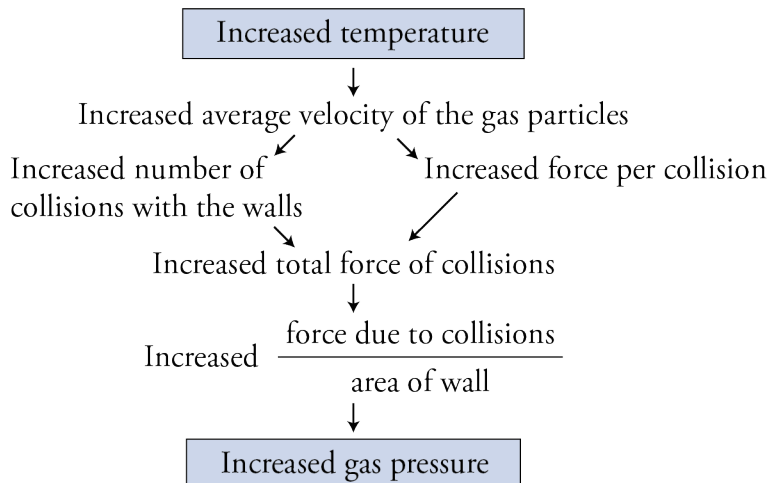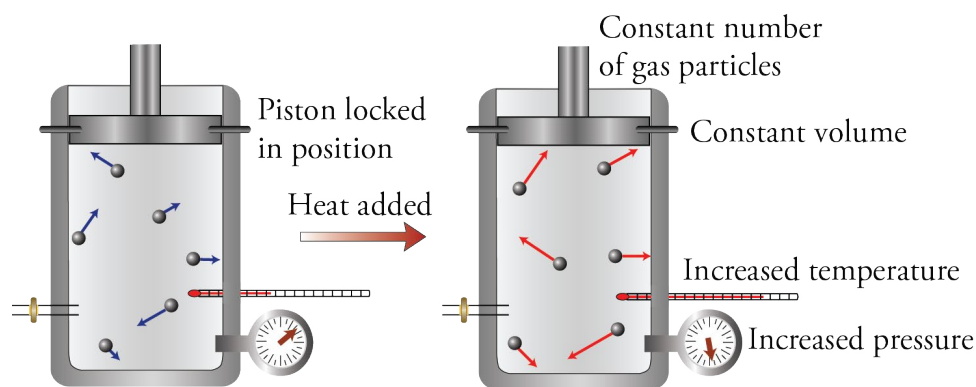
OBJECTIVE 10B

**You can see an animation that demonstrates this relationship at the textbook's Web site.**



Piston locked in position

Heat added

Constant number of gas particles

Constant volume

Increased temperature

Increased pressure

Increased temperature

Increased average velocity of the gas particles

Increased number of collisions with the walls

Increased force per collision

Increased total force of collisions

Increased $\dfrac{\text{force due to collisions}}{\text{area of wall}}$

Increased gas pressure

## The Relationship Between Volume and Temperature

Consider the system shown in Figure 11.5. To demonstrate the relationship between temperature and volume of gas, we keep the number of gas particles and gas pressure constant. If our valve is closed and if our system has no leaks, the number of particles is constant. We keep the gas pressure constant by allowing the piston to move freely throughout our experiment, because then it will adjust to keep the pressure pushing on it from the inside equal to the constant external pressure pushing on it due to the weight of the piston and the atmospheric pressure. The atmospheric pressure is the pressure in the air outside the container, which acts on the top of the piston due to the force of collisions between particles in the air and the top of the piston.

If we increase the temperature, the piston in our apparatus moves up, increasing the volume occupied by the gas. A decrease in temperature leads to a decrease in volume.
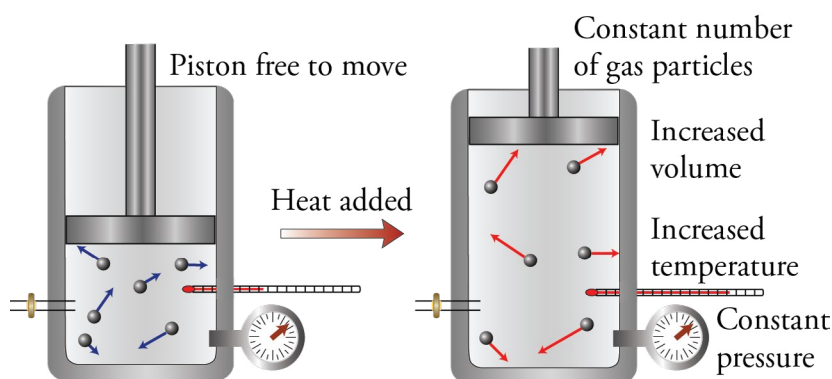
> Increased temperature  →  Increased volume
> Decreased temperature  →  Decreased volume

The increase in temperature of the gas leads to an increase in the average velocity of the gas particles, which leads in turn to more collisions with the walls of the container and a greater force per collision. This greater force acting on the walls of the container leads to an initial increase in the gas pressure. Thus the increased temperature of our gas creates an internal pressure, acting on the bottom of the piston, that is greater than the external pressure. The greater internal pressure causes the piston to move up, increasing the volume of the chamber. The increased volume leads to a decrease in gas pressure in the container, until the internal pressure is once again equal to the constant external pressure (Figure 11.5). Similar reasoning can be used to explain why decreased temperature leads to decreased volume when the number of gas particles and pressure are held constant.

For an ideal gas, *volume and temperature described in kelvins are directly proportional if the number of gas particles and pressure are constant*. This is called **Charles' Law**.
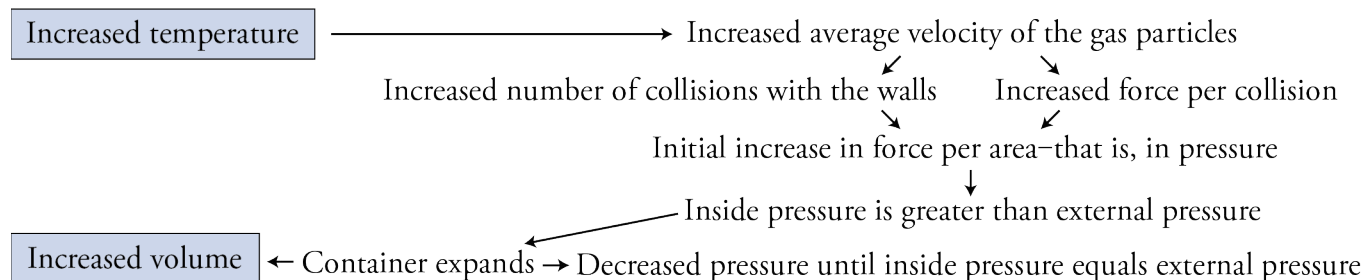
$$V \propto T \quad \text{if n and P are constant}$$

Placing a balloon in liquid nitrogen lowers the temperature of the gas and causes an initial decrease in pressure. With its internal pressure now lower than the pressure of the air outside, the balloon shrinks to a much smaller volume.

The University of Michigan Department of Physics Lecture Demonstration Lab



**Figure 11.5**
**Relationship Between Temperature and Volume** Increased temperature leads to increased volume if the number of gas particles and pressure are constant.

**You can see an animation that demonstrates this relationship at the textbook's Web site.**

Increased temperature ⟶ Increased average velocity of the gas particles

Increased number of collisions with the walls          Increased force per collision

Initial increase in force per area–that is, in pressure

Inside pressure is greater than external pressure

Increased volume  ← Container expands → Decreased pressure until inside pressure equals external pressure

## The Relationship Between Number of Gas Particles and Pressure

OBJECTIVE 10D

To explore the relationship between gas pressure and the number of gas particles, we could set up our experimental system as shown in Figure 11.6. The volume is held constant by locking the piston so it cannot move. The temperature is kept constant by allowing heat to flow in or out of the cylinder in order to keep the temperature of the gas in the cylinder equal to the external temperature. When the number of gas particles is increased by adding gas through the valve on the left of the cylinder, the pressure gauge shows an increase in pressure. When gas is allowed to escape from the valve, the decrease in the number of gas particles causes a decrease in the pressure of the gas.

OBJECTIVE 10D

Increased number of gas particles → Increased pressure

Decreased number of gas particles → Decreased pressure

The increase in the number of gas particles in the container leads to an increase in the number of collisions with the walls per unit time. This leads to an increase in the force per unit area–that is, to an increase in gas pressure.

   If the temperature and the volume of an ideal gas are held constant, the number of gas particles in a container and the gas pressure are directly proportional.

OBJECTIVE 10D

$P \propto n$   if $T$ and $V$ are constant

**Figure 11.6**
**Relationship Between  Number of Gas Particles and Pressure**
Increased number of gas particles leads to increased pressure if the temperature and volume are constant.
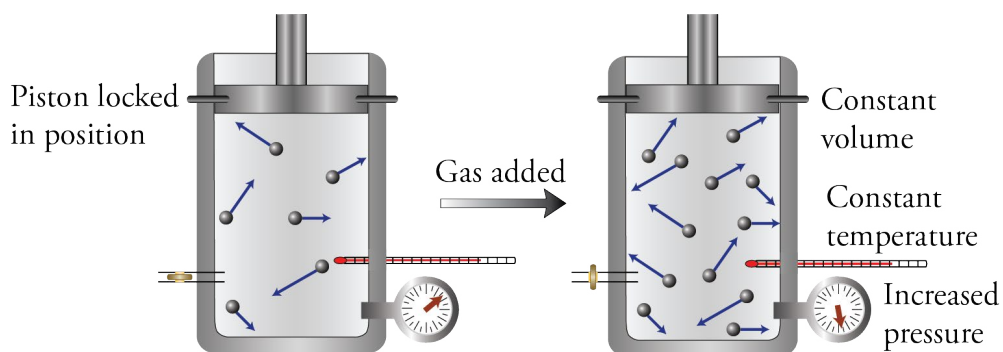
OBJECTIVE 10D

You can see an animation that demonstrates this relationship at the textbook's Web site.



Piston locked in position

Gas added

Constant volume

Constant temperature

Increased pressure

Increased number of gas particles

↓

Increased number of collisions with the walls

↓

Increased total force of collisions

↓

Increased gas pressure
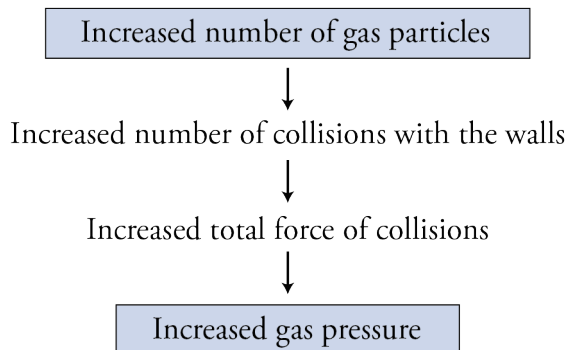
## The Relationship Between Number of Gas Particles and Volume

Figure 11.7 shows how the relationship between the number of gas particles and volume can be demonstrated using our apparatus. Temperature is held constant by allowing heat to move into or out of the system, thus keeping the internal temperature equal to the constant external temperature. Pressure is held constant by allowing the piston to move freely to keep the internal pressure equal to the external pressure. When we increase the number of gas particles in the cylinder by adding gas through the valve on the left of the apparatus, the piston rises, increasing the volume available to the gas. If the gas is allowed to escape from the valve, the volume decreases again.
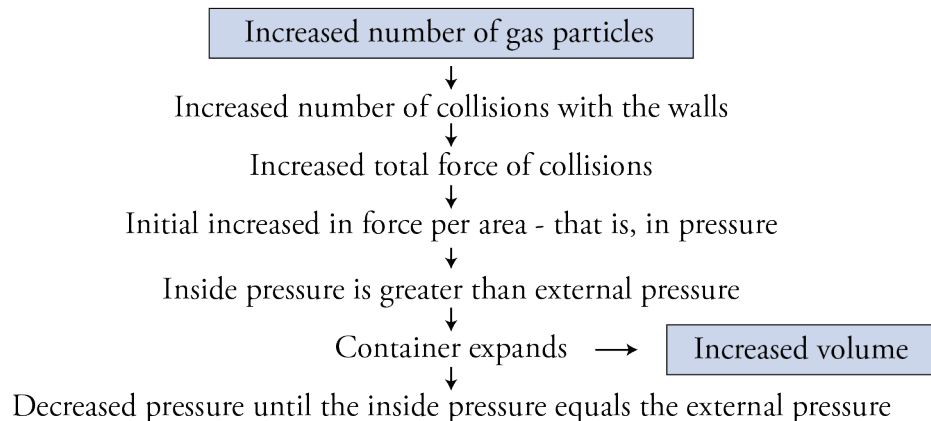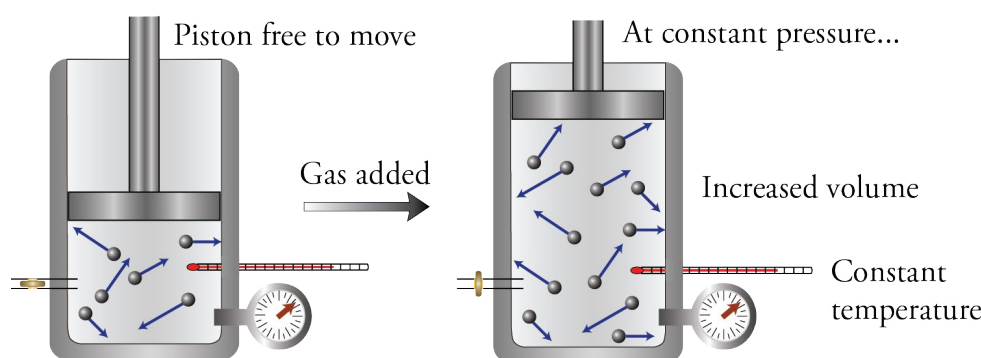
Increased number of gas particles  →  Increased volume
Decreased number of gas particles  →  Decreased volume

The explanation for why an increase in the number of gas particles increases volume starts with the recognition that the increase in the number of gas particles results in more collisions per second against the walls of the container. The greater force due to these collisions creates an initial increase in the force per unit area—or gas pressure—acting on the walls. This will cause the piston to rise, increasing the gas volume and decreasing the pressure until the internal and external pressure are once again equal (Figure 11.7). Take a minute or two to work out a similar series of steps to explain why decreased number of gas particles leads to decreased volume.

The relationship between moles of an ideal gas and volume is summarized by **Avogadro's Law**, which states that *the volume and the number of gas particles are directly proportional if the temperature and pressure are constant.*

$$V \propto n \quad \text{if } T \text{ and } P \text{ are constant}$$



When this person blows more air into the balloon, the increased number of gas particles initially leads to an increased pressure. Because the internal pressure is now greater than the pressure of the air outside the balloon, the balloon expands to a larger volume.



Piston free to move     At constant pressure...

Gas added

Increased volume

Constant temperature

Increased number of gas particles
↓
Increased number of collisions with the walls
↓
Increased total force of collisions
↓
Initial increased in force per area - that is, in pressure
↓
Inside pressure is greater than external pressure
↓
Container expands  →  Increased volume
↓
Decreased pressure until the inside pressure equals the external pressure

**Figure 11.7**
**Relationship Between Number of Gas Particles and Volume**
Increased number of gas particles leads to increased volume if the temperature and pressure are constant.

You can see an animation that demonstrates this relationship at the textbook's Web site.
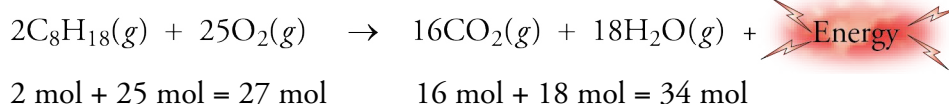
## Gases and the Internal Combustion Engine

Now, let's apply our model for gases to a real situation. Because the changes that take place in a typical car engine illustrate many of the characteristics of gases, let's take a look at how the internal combustion engine works (Figure 11.8).

OBJECTIVE 11     Liquid gasoline is a mixture of hydrocarbons, with from five to twelve carbon atoms in each molecule. It evaporates to form a gas, which is mixed with air and injected into the engine's cylinders. (See cylinder 1 in Figure 11.8.) The movement of the engine's pistons turns a crankshaft that causes the piston in one of the cylinders containing the gasoline air mixture to move up, compressing the gases and increasing the pressure of the gas mixture in the cylinder. (See cylinder 2 in Figure 11.8.)

A spark ignites the mixture of compressed gases in a cylinder, and the hydrocarbon compounds in the gasoline react with the oxygen in the air to form carbon dioxide gas and water vapor. (See cylinder 3 in Figure 11.8.) A typical reaction is
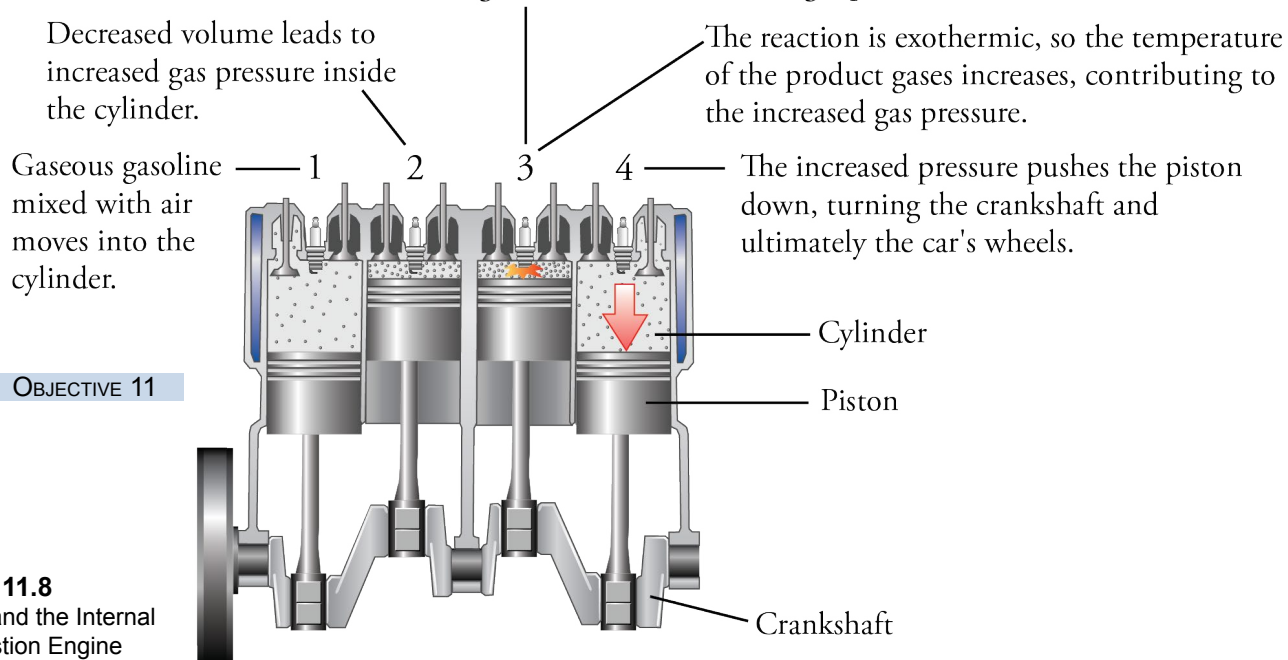
$$2C_8H_{18}(g) \; + \; 25O_2(g) \quad \rightarrow \quad 16CO_2(g) \; + \; 18H_2O(g) \; + \quad Energy$$

2 mol + 25 mol = 27 mol          16 mol + 18 mol = 34 mol

OBJECTIVE 11     In this representative reaction, a total of 27 moles of gas are converted into 34 moles of gas. The increase in moles of gas leads to an increase in number of collisions per second with the walls of the cylinder, which creates greater force acting against the walls and a greater gas pressure in the cylinder. The pressure is increased even more by the increase in the temperature of the gas due to the energy released in the reaction. The increased temperature increases the average velocity of the gas particles, which leads to more frequent collisions with the walls and a greater average force per collision.

OBJECTIVE 11     The increased pressure of the gas pushes the piston down. (See cylinder 4 in Figure 11.8.) This movement of the pistons turns the crankshaft, which, through a series of mechanical connections, turns the wheels of the car.

The combustion of the gasoline leads to an increase in moles of gas, which also causes the gas pressure to increase.

Decreased volume leads to increased gas pressure inside the cylinder.

The reaction is exothermic, so the temperature of the product gases increases, contributing to the increased gas pressure.

Gaseous gasoline mixed with air moves into the cylinder.

1     2     3     4

The increased pressure pushes the piston down, turning the crankshaft and ultimately the car's wheels.

Cylinder

OBJECTIVE 11

Piston

**Figure 11.8**
Gases and the Internal Combustion Engine

Crankshaft

## Explanations for Other Real World Situations

The relationships between gas properties can be used to explain how we breathe in and out. When the muscles of your diaphragm contract, your chest expands, and the volume of your lungs increases. This change leads to a decrease in the number of particles per unit volume inside the lungs, leaving fewer particles near any give area of the inner surface of the lungs. Fewer particles means fewer collisions per second per unit area of lungs and a decrease in force per unit area, or gas pressure. During quiet, normal breathing, this increase in volume decreases the pressure in the lungs to about 0.4 kilopascals lower than the atmospheric pressure. As a result, air moves into the lungs faster than it moves out, bringing in fresh oxygen. When the muscles relax, the lungs return to their original, smaller volume, causing the pressure in the lungs to increase to about 0.4 kilopascals above atmospheric pressure. Air now goes out of the lungs faster than it comes in (Figure 11.9).

OBJECTIVE 12
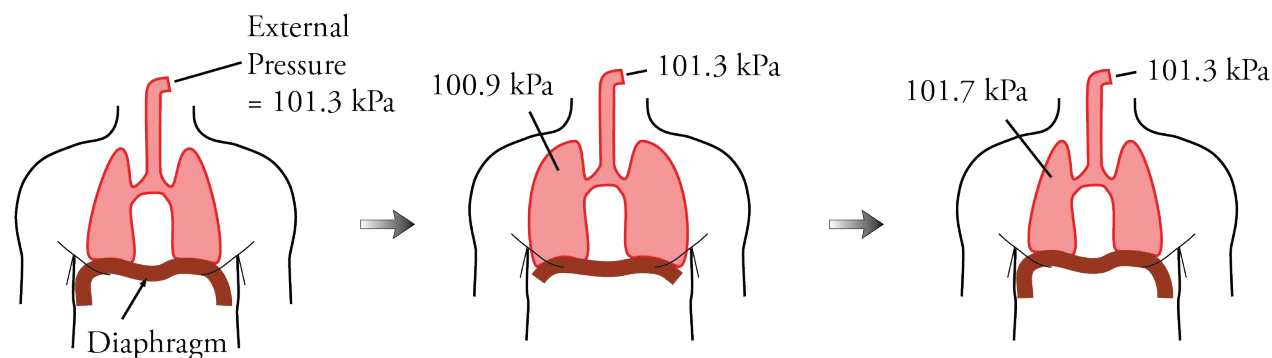
OBJECTIVE 12    **Figure 11.9**
The Mechanics of Breathing

External Pressure = 101.3 kPa

Diaphragm

100.9 kPa          101.3 kPa

The diaphragm contracts, and the chest expands, causing the lungs to expand. The increased volume decreases the pressure in the lungs to below the external pressure, causing air to move into the lungs faster than it moves out.

101.7 kPa          101.3 kPa

The diaphragm relaxes and the chest returns to its original volume, causing the lung volume to decrease. This increases the pressure in the lungs, causing air to move out of the lungs faster than it moves in.

Let's return to the gas-related issues discussed at Lilia's family dinner. When Lilia's brother John adds helium gas to one of his instrument-carrying balloons, its volume increases. Increased number of gas particles leads to increased gas pressure, making the internal pressure greater than the external pressure acting on the outside surface of the balloon. The increase of internal pressure leads to an expansion of the balloon. When the balloon is released and rises into the air, the concentration of gases outside the balloon decreases with increased distance from the earth, leading to a decrease in the atmospheric pressure acting on the outside of the balloon. The greater pressure inside the balloon causes the balloon to expand as it rises. This increase in volume is diminished somewhat by the slight loss of gas from tiny holes in the balloon and by the general decrease in temperature with increased distance from the earth.

When the temperature is increased for the gases in the reaction vessel that Lilia's sister Rebecca is helping design, the gas pressure rises. This is due to an increase in

OBJECTIVE 13

The volume of this balloon will increase as it moves up into the atmosphere.

the average velocity of the particles and therefore an increase in the rate of collisions of gas particles with the constant area of the walls and an increase in the average force per collision. Rebecca's pressure release valve allows gas to escape if the pressure gets to a certain level. The decrease in the number of gas particles when gas escapes from the valve keeps the pressure below dangerous levels.

To answer Ted's question about how to calculate gas density and to see how Amelia can estimate the number of tubes she can fill for her neon light sculptures, we need to continue on to the next section.

## 11.2  Ideal Gas Calculations

This section shows how to do calculations such as those necessary to answer Ted's and Amelia's questions about gas density and volume, and in addition, it considers some of the gas-related issues that Lilia's sister Rebecca and her co-workers need to resolve. For one thing, the design team needs to know the amount of gas that they can safely add to their reaction vessel, and then Rebecca needs to determine the maximum temperature at which the reaction can be run without causing the pressure of that amount of gas to reach dangerous levels.

All these calculations, and others like them, can be done with the aid of two useful equations that we will now derive from the relationships described in Section 11.1.

### Calculations Using the Ideal Gas Equation

We discovered in Section 11.1 that pressure of an ideal gas is directly proportional to the number of gas particles (expressed in moles), directly proportional to temperature, and inversely proportional to the volume of the container.

$$P \propto n \quad \text{if } T \text{ and } V \text{ are constant}$$
$$P \propto T \quad \text{if } n \text{ and } V \text{ are constant}$$
$$P \propto \frac{1}{V} \quad \text{if } n \text{ and } T \text{ are constant}$$

These three relationships can be summarized in a single equation:

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

Another way to express the same relationship is

$$P = (\text{a constant}) \frac{nT}{V}$$

The constant in this equation is the same for all ideal gases. It is called the **universal gas constant** and is expressed with the symbol **R**. The value of R depends on the units of measure one wishes to use in a given calculation. Two choices are given below, showing R for different pressure units (atmospheres, atm, and kilopascals, kPa).

OBJECTIVE 14

$$R = \frac{0.082058 \text{ L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \quad \text{or} \quad \frac{8.3145 \text{ L} \cdot \text{kPa}}{\text{K} \cdot \text{mol}}$$

Substituting $R$ for "a constant" and rearranging the equation yields the *ideal gas equation* (often called the ideal gas law) in the form that is most commonly memorized and written:

$$PV = nRT$$

Because we will often be interested in the masses of gas samples, it is useful to remember an expanded form of the ideal gas equation that uses mass in grams ($g$) divided by molar mass ($M$) instead of moles ($n$).

$$n = \text{moles} = \frac{\text{grams}}{\dfrac{\text{grams}}{\text{mole}}} = \frac{\text{mass in grams}}{\text{molar mass}} = \frac{g}{M}$$

$$PV = \frac{g}{M}RT$$

The following sample study sheet describes how calculations can be done using the two forms of the ideal gas equation.

---

**TIP-OFF**   The usual tip-off that you can use the ideal gas equation to answer a question is that you are given three properties of a sample of gas and asked to calculate the fourth. A more general tip-off is that only one gas is mentioned and there are no changing properties.

**GENERAL STEPS**   Follow these steps.

**STEP 1**   Assign variables to the values given and the value that is unknown. Use $P$ for pressure, $V$ for volume, $n$ for moles, $T$ for temperature, $g$ for mass, and $M$ for molar mass.

**STEP 2**   Write the appropriate form of the ideal gas equation.
  ■ If the number of particles in moles is given or desired, use the most common form of the ideal gas equation.

$$PV = nRT \qquad R = \frac{0.082058 \text{ L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \quad \text{or} \quad \frac{8.3145 \text{ L} \cdot \text{kPa}}{\text{K} \cdot \text{mol}}$$

  ■ If mass or molar mass is given or desired, use the expanded form of the ideal gas equation.

$$PV = \frac{g}{M}RT \quad g = \text{mass} \quad M = \text{molar mass}$$

**STEP 3**   Rearrange the equation to isolate the unknown.

**STEP 4**   Plug in the known values, including units. Be sure to use Kelvin temperatures.

**STEP 5**   Make any necessary unit conversions and cancel your units.

**STEP 6**   Calculate your answer and report it to the correct significant figures and with the correct unit.

**EXAMPLE**   See Examples 11.1 and 11.2.

Sample Study Sheet 11.1

Using the Ideal Gas Equation

## EXAMPLE 11.1 - Using the Ideal Gas Equation

OBJECTIVE 15

Incandescent light bulbs "burn out" because their tungsten filament evaporates, weakening the thin wire until it breaks. Argon gas is added inside the bulbs to reduce the rate of evaporation. (Argon is chosen because, as a noble gas, it will not react with the components of the bulb, and because it is easy to obtain in significant quantities. It is the third most abundant element in air.) What is the pressure in atmospheres of $3.4 \times 10^{-3}$ moles of argon gas in a 75-mL incandescent light bulb at 20 °C?

*Solution*

We are given three of the four properties of ideal gases (moles, volume, and temperature), and we are asked to calculate the fourth (pressure). Therefore, we use the ideal gas equation to solve this problem.

**STEP 1** We assign variables to the values that we are given and to the unknown value. Remember to use Kelvin temperatures in gas calculations.

$$P = ? \qquad n = 3.4 \times 10^{-3} \text{ mol} \qquad V = 75 \text{ mL} \qquad T = 20 \text{ °C} + 273.15 = 293 \text{ K}$$

**STEP 2** We pick the appropriate form of the ideal gas equation. Because moles of gas are mentioned in the problem, we use

$$PV = nRT$$

**STEP 3** We rearrange our equation to isolate the unknown property.

$$P = \frac{nRT}{V}$$

**STEP 4** We plug in the values that are given, including their units. We use the value for $R$ that contains the unit we want to calculate, atm.

$$P = \frac{3.4 \times 10^{-3} \text{ mol} \left( \dfrac{0.082058 \text{ L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \right) 293 \text{ K}}{75 \text{ mL}}$$

**STEPS 5 AND 6** If the units do not cancel to yield the desired unit, we do the necessary unit conversions to make them cancel. In this example, we need to convert 75 mL to L so that the volume units cancel. We finish the problem by calculating the answer and reporting it with the correct significant figures and with the correct unit.

$$P = \frac{3.4 \times 10^{-3} \text{ mol} \left( \dfrac{0.082058 \text{ L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \right) 293 \text{ K}}{75 \text{ mL}} \left( \frac{10^3 \text{ mL}}{1 \text{ L}} \right) = \textbf{1.1 atm}$$

To see a practical application of calculations like Example 11.1, let's take a closer look at one of the gas-related issues that Lilia's sister Rebecca and her coworkers are dealing with. The apparatus Rebecca is helping to design will be used in the first step of the commercial process that makes nitric acid for fertilizers. In this step, the gases ammonia and oxygen are converted into gaseous nitrogen monoxide and water.

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$$
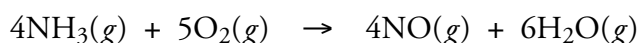
When the team of chemists and chemical engineers meets to design the new process, they begin with some preliminary calculations. First, they want to determine the amount of gas that they can safely add to their reaction vessel. They know that the optimum conditions for this reaction are a temperature of about 825 °C and a pressure of about 700 kPa, and they have been told by the plant architect that the maximum volume of the reaction vessel will be 2500 m$^3$. Notice that the value for $R$ that contains the unit kPa is used.

$$PV = nRT$$

$$n = \frac{PV}{RT} = \frac{700 \text{ k\bcancel{Pa}} (2500 \text{ \bcancel{m}}^3)}{\left(\dfrac{8.3145 \text{ \bcancel{L}} \cdot \text{k\bcancel{Pa}}}{\text{\bcancel{K}} \cdot \text{mol}}\right)(825 + 273.15)\text{\bcancel{K}}}\left(\frac{10^3 \text{ \bcancel{L}}}{1 \text{ \bcancel{m}}^3}\right) = 1.92 \times 10^5 \text{ mol}$$

As you practice using the ideal gas equation, you may be tempted to save time by plugging in the numbers without including their accompanying units. Be advised, however, that it is always a good idea to include the units as well. If the units cancel to yield a reasonable unit for the unknown property, you can feel confident that

- You picked the correct equation.
- You did the algebra correctly to solve for your unknown.
- You have made the necessary unit conversions.

If the units do not cancel to yield the desired unit, check to be sure that you have done the algebra correctly and that you have made the necessary unit conversions.

## EXAMPLE 11.2 - Using the Ideal Gas Equation

Incandescent light bulbs were described in Example 11.1. At what temperature will 0.0421 g of Ar in a 23.0-mL incandescent light bulb have a pressure of 952 mmHg?

OBJECTIVE 16

*Solution*

$$T = ? \qquad g = 0.0421 \text{ g} \qquad V = 23.0 \text{ mL} \qquad P = 952 \text{ mmHg}$$

The tip off that this is an ideal gas equation problem is that we are given three properties of gases and asked to calculate the fourth.

Because the gas's mass is given, we choose the expanded form of the ideal gas equation. We rearrange the equation to isolate the unknown volume, plug in given values, and cancel our units. In order to cancel the pressure and volume units, we must convert 952 mmHg into atmospheres and convert 23.0 mL into liters. When this is done, the units cancel to yield an answer in kelvins. Because kelvin is a reasonable temperature unit, we can assume that we have picked the correct equation, done the algebraic manipulation correctly, and made all of the necessary unit conversions.

$$PV = \frac{g}{M}RT \qquad g = \text{mass} \qquad M = \text{molar mass}$$

$$T = \frac{PVM}{gR} = \frac{952 \text{ \bcancel{mmHg}} (23.0 \text{ \bcancel{mL}})\, 39.948 \dfrac{\bcancel{g}}{\text{mol}}}{0.0421 \text{\bcancel{g}} \left(\dfrac{0.082058 \text{ \bcancel{L}} \cdot \text{\bcancel{atm}}}{\text{K} \cdot \text{\bcancel{mol}}}\right)}\left(\frac{1 \text{ \bcancel{atm}}}{760 \text{ \bcancel{mmHg}}}\right)\left(\frac{1 \text{ \bcancel{L}}}{10^3 \text{ \bcancel{mL}}}\right)$$

$$= 333 \text{ K or } 60 \text{ °C}$$

## EXAMPLE 11.3 - Using the Ideal Gas Equation

OBJECTIVE 17

In the incandescent light bulb described in Example 11.1, what is the density of argon gas at 80.8 °C and 131 kPa?

*Solution*

Because only one gas is mentioned and because there are no changing properties, we recognize this as an ideal gas equation problem.

Density is mass divided by volume. Using variables from the expanded form of the ideal gas equation, it can be expressed as $g/V$.

$$\frac{g}{V} = ? \qquad P = 131 \text{ kPa} \qquad T = 80.8 \text{ °C} + 273.15 = 354.0 \text{ K}$$

$$PV = \frac{g}{M} RT$$

To save a little time, we can use the value for $R$ that contains the pressure unit given in the problem, kPa.

$$\frac{g}{V} = \frac{PM}{RT} = \frac{131 \text{ kPa} \left( \dfrac{39.948 \text{ g}}{1 \text{ mol}} \right)}{\left( \dfrac{8.3145 \text{ L} \cdot \text{kPa}}{\text{K} \cdot \text{mol}} \right) (354.0 \text{ K})} = \mathbf{1.78 \text{ g/L}}$$

Alternatively, we could use the same value for $R$ that we used in Examples 11.1 and 11.2, along with a unit conversion.

$$\frac{g}{V} = \frac{PM}{RT} = \frac{131 \text{ kPa} \left( \dfrac{39.948 \text{ g}}{1 \text{ mol}} \right)}{\left( \dfrac{0.082058 \text{ L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \right) (354.0 \text{ K})} \left( \frac{1 \text{ atm}}{101.325 \text{ kPa}} \right) = \mathbf{1.78 \text{ g/L}}$$

A disastrous series of crashes in the 1992 Indianapolis 500 race has been traced to an unfortunate combination of factors, including an unusually and unexpectedly low temperature, which led to a higher-than-expected air density that day. A change in air density causes changes in aerodynamic forces, engine power, and brake performance.

If such a change is anticipated, car designers and mechanics, such as Lilia's brother Ted, can make adjustments. For example, because of its altitude, Denver has about 15% lower air density than does a city at sea level. Therefore, Indy cars that race in Denver are fitted with larger-than-normal brakes to compensate for the lower air resistance and greater difficulty stopping. Ted can do calculations similar to Example 11.3 to determine the density of air on different days and at different racetracks. An example is the following set of equations calculating the density of air at 762 mmHg and 32 °C and at 758 mmHg and 4 °C. (The average molar mass for the gases in air is 29 g/mol.)

$$PV = \frac{g}{M}RT$$

OBJECTIVE 17

$$\frac{g}{V} = \frac{PM}{RT} = \frac{762 \text{ mmHg} \left(\dfrac{29 \text{ g}}{1 \text{ mol}}\right)}{\left(\dfrac{0.082058 \text{ L} \cdot \text{atm}}{\text{K} \cdot \text{mol}}\right) 305 \text{ K}} \left(\frac{1 \text{ atm}}{760 \text{ mmHg}}\right) = 1.2 \text{ g/L}$$

$$\frac{g}{V} = \frac{PM}{RT} = \frac{758 \text{ mmHg} \left(\dfrac{29 \text{ g}}{1 \text{ mol}}\right)}{\left(\dfrac{0.082058 \text{ L} \cdot \text{atm}}{\text{K} \cdot \text{mol}}\right) 277 \text{ K}} \left(\frac{1 \text{ atm}}{760 \text{ mmHg}}\right) = 1.3 \text{ g/L}$$

You can read about why the measured, real properties differ from the properties calculated from the ideal gas equation at the textbook's Web site.

## EXERCISE 11.1 - Using the Ideal Gas Equation

Krypton gas does a better job than argon of slowing the evaporation of the tungsten filament in an incandescent light bulb. Because of its higher cost, however, krypton is only used when longer life is considered to be worth the extra expense.

OBJECTIVE 15
OBJECTIVE 16
OBJECTIVE 17

a.  How many moles of krypton gas must be added to a 175-mL incandescent light bulb to yield a gas pressure of 117 kPa at 21.6 °C?

b.  What is the volume of an incandescent light bulb that contains 1.196 g Kr at a pressure of 1.70 atm and a temperature of 97 °C?

c.  What is the density of krypton gas at 18.2 °C and 762 mmHg?

## When Properties Change

Another useful equation, derived from the ideal gas equation, can be used to calculate changes in the properties of a gas. In the first step of the derivation, the ideal gas equation is rearranged to isolate $R$.

$$\frac{PV}{nT} = R$$

In this form, the equation shows that no matter how the pressure, volume, moles of gas, or temperature of an ideal gas may change, the other properties will adjust so that the ratio of $PV/nT$ always remains the same. In the equations below, $P_1$, $V_1$, $n_1$, and $T_1$ are the initial pressure, volume, moles of gas, and temperature, and $P_2$, $V_2$, $n_2$, and $T_2$ are the final pressure, volume, moles of gas, and temperature.

$$P_1 V_1 = n_1 R T_1 \quad \text{so} \quad \frac{P_1 V_1}{n_1 T_1} = R$$

$$P_2 V_2 = n_2 R T_2 \quad \text{so} \quad \frac{P_2 V_2}{n_2 T_2} = R$$

Therefore,

OBJECTIVE 18

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$$

The equation above is often called the *combined gas law equation*. It can be used to do calculations such as Example 11.4 and Exercise 11.2, which follow the sample study sheet below.

---

**Sample Study Sheet 11.2**
**Using the Combined Gas Law Equation**

OBJECTIVE 18

**TIP-OFF** – The problem requires calculating a value for a gas property that has changed. In other words, you are asked to calculate a new pressure, temperature, moles (or mass), or volume of gas given sufficient information about the initial and other final properties.

**GENERAL STEPS**

**STEP 1**   Assign the variables $P$, $T$, $n$, and $V$ to the values you are given and to the unknown value. Use the subscripts 1 and 2 to show initial or final conditions.

**STEP 2**   Write out the combined gas law equation, but eliminate the variables for any constant properties. (You can assume that the properties not mentioned in the problem remain constant.)

**STEP 3**   Rearrange the equation to isolate the unknown property.

**STEP 4**   Plug in the values for the given properties.

**STEP 5**   Make any necessary unit conversions and cancel your units.

**STEP 6**   Calculate your answer and report it with the correct units and with the correct significant figures.

**EXAMPLE**   See Example 11.4.

## EXAMPLE 11.4 - Using the Combined Gas Law Equation

Neon gas in luminous tubes radiates red light—the original "neon light." The standard gas containers used to fill the tubes have a volume of 1.0 L and store neon gas at a pressure of 101 kPa at 22 °C. A typical luminous neon tube contains enough neon gas to exert a pressure of 1.3 kPa at 19 °C. If all the gas from a standard container is allowed to expand until it exerts a pressure of 1.3 kPa at 19 °C, what will its final volume be? If Lilia's sister Amelia is adding this gas to luminous tubes that have an average volume of 500 mL, what is the approximate number of tubes she can fill?

Neon tubes

*Solution*

We recognize this as a combined gas law problem because it requires calculating a value for a gas property that has changed. In this case, that property is volume.

**STEP 1**  We assign variables to the given values and to the unknown.

$V_1 = 1.0$ L    $P_1 = 101$ kPa        $T_1 = 22$ °C $+ 273.15 = 295$ K

$V_2 = ?$        $P_2 = 1.3$ kPa        $T_2 = 19$ °C $+ 273.15 = 292$ K

**STEP 2**  We write the combined gas law equation, eliminating variables for properties that are constant. Because moles of gas are not mentioned, we assume that they are constant ($n_1 = n_2$).

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \quad \text{becomes} \quad \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

**STEPS 3 AND 4**  We rearrange the equation to solve for our unknown and plug in the given values.

$$V_2 = V_1 \left(\frac{T_2}{T_1}\right)\left(\frac{P_1}{P_2}\right) = 1.0 \text{ L} \left(\frac{292 \text{ K}}{295 \text{ K}}\right)\left(\frac{101 \text{ kPa}}{1.3 \text{ kPa}}\right)$$

**STEP 5**  Our units cancel to yield liters, which is a reasonable volume unit, so we do not need to make any unit conversions.

**STEP 6**  We finish the problem by calculating our answer and reporting it with the correct significant figures.

$$V_2 = V_1 \left(\frac{T_2}{T_1}\right)\left(\frac{P_1}{P_2}\right) = 1.0 \text{ L} \left(\frac{292 \text{ K}}{295 \text{ K}}\right)\left(\frac{101 \text{ kPa}}{1.3 \text{ kPa}}\right) = 77 \text{ L}$$

We can now answer Amelia's question about how many neon tubes she can expect to fill from each neon cylinder.

$$? \text{ tubes} = 77 \text{ L} \left(\frac{10^3 \text{ mL}}{1 \text{ L}}\right)\left(\frac{1 \text{ tube}}{500 \text{ mL}}\right) = 1.5 \times 10^2 \text{ tubes, or about 150 tubes}$$

Let's look at another gas-related issue that Lilia's sister Rebecca needs to consider in designing the pressure valve for the reaction vessel at her chemical plant. She knows that, for safety reasons, the overall pressure must be kept below 1000 kPa, and she knows that the most likely cause of increased pressure is increased temperature. To get an idea of how high the temperature can go safely, she could use the combined gas law equation to calculate the temperature at which the pressure of the gas will reach 1000 kPa if the initial temperature was 825 °C (1098 K) and the initial pressure was
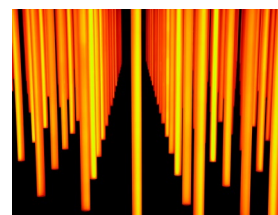
700.0 kPa. The reaction vessel has a constant volume, and we can assume that it has no leaks. Therefore, the volume and moles of gas remain constant.

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \quad \text{or} \quad \frac{P_1}{T_1} = \frac{P_2}{T_2} \quad \text{because } V_1 = V_2 \text{ and } n_1 = n_2$$

$$T_2 = T_1 \left(\frac{P_2}{P_1}\right) = 1098 \text{ K} \left(\frac{1000 \text{ kPa}}{700.0 \text{ kPa}}\right) = \textbf{1569 K or 1295 °C}$$

Although Rebecca designs her pressure release valve to allow gas to escape when the pressure reaches 1000 kPa, she hopes it will never need to be used. To avoid the release of gases, she informs the rest of her team that they should be careful to design the process to keep the temperature well below 1295 °C.

## EXERCISE 11.2 - Using the Combined Gas Law Equation

OBJECTIVE 18

A helium weather balloon is filled in Monterey, California, on a day when the atmospheric pressure is 102 kPa and the temperature is 18 °C. Its volume under these conditions is $1.6 \times 10^4$ L. Upon being released, it rises to an altitude where the temperature is −8.6 °C, and its volume increases to $4.7 \times 10^4$ L. Assuming that the internal pressure of the balloon equals the atmospheric pressure, what is the pressure at this altitude?

## 11.3  Equation Stoichiometry and Ideal Gases

Chemists commonly want to convert from the amount of one substance in a chemical reaction to an amount of another substance in that reaction. For example, they might want to calculate the maximum amount of a product that can be formed from a certain amount of a reactant, or they might want to calculate the minimum amount of one reactant that would be necessary to use up a known amount of another reactant. Chapter 10 showed how to do conversions like these. In Example 10.1, we calculated the minimum mass of water in kilograms necessary to react with $2.50 \times 10^4$ kg of tetraphosphorus decoxide, $P_4O_{10}$, in the reaction

$$P_4O_{10}(s) + 6H_2O(l) \quad \rightarrow \quad 4H_3PO_4(aq)$$
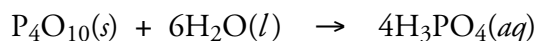
This kind of problem, an equation stoichiometry problem, is generally solved in the following steps:

Measurable property of 1 → moles 1 → moles 2 → measurable property of 2

For substances that are pure solids and liquids, such as $P_4O_{10}$ and $H_2O$, mass is an easily measured property, so the specific steps used in solving Example 10.1 were

Mass $P_4O_{10}$  →  moles $P_4O_{10}$  →  moles $H_2O$  →  mass $H_2O$

Amounts of gas, however, are more commonly described in terms of volume at a particular temperature and pressure. (It is far easier to measure a gas's volume than to measure its mass.) Therefore, to solve equation stoichiometry problems involving gases, we must be able to convert between the volume of a gaseous substance and the corresponding number of moles. For example, if both substances involved in the equation stoichiometry problem are gases, the general steps would be

Volume of gas 1   →   moles 1   →   moles 2   →   volume of gas 2

There are several ways to convert between moles of a gaseous substance and its volume. One approach is to use as a conversion factor the molar volume at STP. STP stands for "standard temperature and pressure." *Standard temperature* is 0 °C, or 273.15 K, and *standard pressure* is 1 atm, or 101.325 kPa, or 760 mmHg. The molar volume, or liters per mole of an ideal gas, at STP can be calculated from the ideal gas equation.

$$PV = nRT$$

$$\frac{V}{n} = \frac{RT}{P} = \frac{\left(\dfrac{8.3145 \text{ L} \cdot \cancel{\text{kPa}}}{\cancel{\text{K}} \cdot \text{mol}}\right)(273.15 \ \cancel{\text{K}})}{101.325 \ \cancel{\text{kPa}}} = \left(\frac{22.414 \text{ L}}{1 \text{ mol}}\right)_{\text{STP}}$$

Because the ideal gas equation applies to all ideal gases, the molar volume at STP applies to all gases that exhibit the characteristics of the ideal gas model. In equation stoichiometry, the molar volume at STP is used in much the way we use molar mass. Molar mass converts between moles and the measurable property of mass; molar volume at STP converts between moles and the measurable property of volume of gas. Note that while every substance has a different molar mass, all ideal gases have the same molar volume at STP. Example 11.5 provides a demonstration.

## Example 11.5 - Gas Stoichiometry

How many liters of carbon dioxide at STP will be formed from the complete combustion of 82.60 g of ethanol, $C_2H_5OH(l)$?

Objective 19

*Solution*

Because we are converting from units of one substance to units of another substance, both involved in a chemical reaction, we recognize this problem as an equation stoichiometry problem. We can therefore use unit analysis for our calculation.

Equation stoichiometry problems have at their core the conversion of moles of one substance to moles of another substance. The conversion factor that accomplishes this part of the calculation comes from the coefficients in the balanced equation. Although we have not been given the balanced equation for the combustion of ethanol, we can supply it ourselves by remembering that when a hydrocarbon compound burns completely, all its carbon forms carbon dioxide, and all its hydrogen forms water.

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$
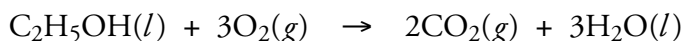
First, however, we convert from mass of $C_2H_5OH$ to moles, using the compound's molar mass. Then we set up the mole-to-mole conversion, using the molar ratio of ethanol to carbon dioxide derived from the coefficients in the balanced equation. The next step is the new one; we use the molar volume at STP to convert from moles of $CO_2$ to volume of $CO_2$ at STP. The sequence as a whole is

Mass $C_2H_5OH \rightarrow$ moles $C_2H_5OH \rightarrow$ moles $CO_2 \rightarrow$ volume of $CO_2$ at STP

$$? \text{ L } CO_2 = 82.60 \text{ g } \cancel{C_2H_5OH} \left(\frac{1 \text{ mol } \cancel{C_2H_5OH}}{46.0692 \text{ g } \cancel{C_2H_5OH}}\right)\left(\frac{2 \text{ mol } \cancel{CO_2}}{1 \text{ mol } \cancel{C_2H_5OH}}\right)\left(\frac{22.414 \text{ L } CO_2}{1 \text{ mol } \cancel{CO_2}}\right)_{\text{STP}}$$

$$= 80.37 \text{ L } CO_2$$

The molar volume at STP—22.414 L/mol—is useful for conversions as long as the gases are at 0 °C and 1 atm of pressure. When the temperature or pressure changes,

the volume of a mole of gas will also change, so 22.414 L/mole cannot be used as a conversion factor for conditions other than STP. One way to convert between volume of gas and moles of gas at temperatures and pressures other than 0 °C and 1 atm is to use the ideal gas equation. We convert from volume to moles by solving the ideal gas equation for $n$ and plugging in the given values for $V$, $T$, and $P$:

$$PV = nRT \quad \text{leads to} \quad n = \frac{PV}{RT}$$

We convert from moles to volume by solving the ideal gas equation for $V$ and plugging in the given values for $n$, $T$, and $P$:

$$PV = nRT \quad \text{leads to} \quad V = \frac{nRT}{P}$$

This technique is demonstrated in Example 11.6.

## EXAMPLE 11.6 - Gas Stoichiometry for Conditions Other Than STP

Ammonia is produced when nitrogen and hydrogen gases react at high pressures and temperatures:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$
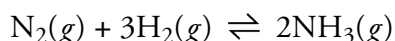
At intervals, the system is cooled to between −10 °C and −20 °C, causing some of the ammonia to liquefy so that it can be separated from the remaining nitrogen and hydrogen gases. The gases are then recycled to make more ammonia. An average ammonia plant might make 1000 metric tons of ammonia per day. When $4.0 \times 10^7$ L of hydrogen gas at 503 °C and 155 atm reacts with an excess of nitrogen, what is the maximum volume of gaseous ammonia that can be formed at 20.6 °C and 1.007 atm?

*Solution*

This equation stoichiometry problem can be solved in three steps:

Volume of $H_2(g)$ $\rightarrow$ moles $H_2$ $\rightarrow$ moles $NH_3$ $\rightarrow$ volume of $NH_3(g)$
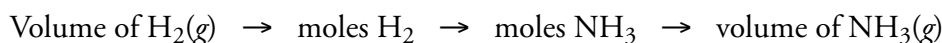
The first step can be accomplished using the ideal gas equation to convert from volume of $H_2$ to moles of $H_2$ under the initial conditions.

$$PV = nRT$$

$$n_{H_2} = \frac{PV}{RT} = \frac{155 \text{ atm} (4.0 \times 10^7 \text{ L})}{\left(\dfrac{0.082058 \text{ L·atm}}{\text{K·mol}}\right)(776 \text{ K})} = 9.7 \times 10^7 \text{ mol } H_2$$

The second step is the typical equation stoichiometry mole-to-mole conversion, using the coefficients from the balanced equation:

$$? \text{ mol } NH_3 = 9.7 \times 10^7 \text{ mol } H_2 \left(\frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2}\right) = 6.5 \times 10^7 \text{ mol } NH_3$$

In the third step, we use the ideal gas equation to convert from moles of $NH_3$ to volume of $NH_3(g)$ under the final conditions:

$$V_{NH_3} = \frac{nRT}{P} = \frac{6.5 \times 10^7 \text{ mol} \left(\dfrac{0.082058 \text{ L·atm}}{\text{K·mol}}\right) 293.8 \text{ K}}{1.007 \text{ atm}} = 1.6 \times 10^9 \text{ L } NH_3$$

An alternative technique allows us to work Example 11.6 and problems like it using a single unit analysis setup. This technique, illustrated in Example 11.7, uses the universal gas constant, $R$, as a conversion factor.

## EXAMPLE 11.7 - Gas Stoichiometry for Conditions Other Than STP (alternative technique)

The question is the same as in Example 11.6: When $4.0 \times 10^7$ L of hydrogen gas at 503 °C and 155 atm reacts with an excess of nitrogen, what is the maximum volume of gaseous ammonia that can be formed at 20.6 °C and 1.007 atm?

*Solution*

We can use $R$ as a conversion factor to convert from volume in liters of $H_2$ to moles of $H_2$. We use it in the inverted form, with liters (L) on the bottom, so that liters will be canceled out.

$$? \text{ L } NH_3 = 4.0 \times 10^7 \text{ L } H_2 \left( \frac{K \cdot mol}{0.082058 \text{ L} \cdot atm} \right)$$

The universal gas constant $(R)$ is different from other conversion factors in that it contains four units rather than two. When we use it to convert from liters to moles, its presence introduces the units of atm and K. We can cancel these units, however, with a ratio constructed from the temperature and pressure values that we are given for $H_2$:

$$? \text{ L } NH_3 = 4.0 \times 10^7 \text{ L } H_2 \left( \frac{K \cdot mol}{0.082058 \text{ L} \cdot atm} \right) \left( \frac{155 \text{ atm}}{776 \text{ K}} \right)$$

We now insert a factor to convert from moles of $H_2$ to moles of $NH_3$, using the coefficients from the balanced equation.

$$? \text{ L } NH_3 = 4.0 \times 10^7 \text{ L } H_2 \left( \frac{K \cdot mol}{0.082058 \text{ L} \cdot atm} \right) \left( \frac{155 \text{ atm}}{776 \text{ K}} \right) \left( \frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2} \right)$$

We complete the sequence by inserting $R$ once again, this time to convert moles of $NH_3$ to volume of $NH_3$ in liters. We eliminate the K and atm units by using a ratio constructed from the temperature and pressure of $NH_3$.

$$? \text{ L } NH_3 = 4.0 \times 10^7 \text{ L } H_2 \left( \frac{K \cdot mol}{0.082058 \text{ L} \cdot atm} \right) \left( \frac{155 \text{ atm}}{776 \text{ K}} \right) \left( \frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2} \right) \left( \frac{0.082058 \text{ L} \cdot atm}{K \cdot mol} \right) \left( \frac{293.8 \text{ K}}{1.007 \text{ atm}} \right)$$

$$= 1.6 \times 10^9 \text{ L } NH_3$$

> The textbook's Web site has a section that describes a shortcut for working equation stoichiometry problems in which volume of one gas is converted into volume of another gas at the same temperature and pressure.

Between this chapter and Chapter 10, we have now seen three different ways to convert between a measurable property and moles in equation stoichiometry problems. The different paths are summarized in Figure 11.10 in the sample study sheet on the next two pages. For pure liquids and solids, we can convert between mass and moles, using the molar mass as a conversion factor. For gases, we can convert between volume of gas and moles using the methods described above. For solutions, molarity provides a conversion factor that enables us to convert between moles of solute and volume of solution. Equation stoichiometry problems can contain any combination of two of these conversions, such as we see in Example 11.8.

## EXAMPLE 11.8 - Equation Stoichiometry

OBJECTIVE 19

The gastric glands of the stomach secrete enough hydrochloric acid to give the stomach juices a hydrochloric acid concentration of over 0.01 M. Some of this stomach acid can be neutralized by an antacid that contains calcium carbonate. The $CaCO_3$ reacts with HCl in the stomach to form aqueous calcium chloride, carbon dioxide gas, and water. A regular antacid tablet will react with about 80 mL of 0.050 M HCl. How many milliliters of $CO_2$ gas at 37 °C and 1.02 atm will form from the complete reaction of 80 mL of 0.050 M HCl?

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

*Solution*

The general steps for this conversion are

Volume of HCl solution → moles HCl → moles $CO_2$ → volume of $CO_2$ gas

$$? \text{ mL } CO_2 = 80 \text{ mL HCl soln} \left( \frac{0.050 \text{ mol HCl}}{10^3 \text{ mL HCl soln}} \right) \left( \frac{1 \text{ mol } CO_2}{2 \text{ mol HCl}} \right) \left( \frac{0.082058 \text{ L} \cdot \text{atm}}{K \cdot \text{mol}} \right) \left( \frac{310 \text{ K}}{1.02 \text{ atm}} \right) \left( \frac{10^3 \text{ mL}}{1 \text{ L}} \right)$$

$$= 50 \text{ mL } CO_2$$

The following sample study sheet summarizes the methods we have learned for solving equation stoichiometry problems.

**Tip-off** – You are asked to convert an amount of one substance to an amount of another substance, both involved in a chemical reaction.

**General Steps**

**Step 1**  Convert the amount of substance 1 to moles 1.

For a pure solid or liquid, this will be a molar mass conversion from mass to moles.

$$\underline{\quad\quad} \; \cancel{g\;1} \left( \frac{1 \text{ mol } 1}{(\text{molar mass}) \; \cancel{g\;1}} \right)$$

For solutions, molarity can be used to convert from volume of solution to moles.

$$\underline{\quad\quad} \; \cancel{\text{L solution } 1} \left( \frac{(\text{number from molarity}) \text{ mol } 1}{1 \; \cancel{\text{L solution } 1}} \right)$$

or

$$\underline{\quad\quad} \; \cancel{\text{mL solution } 1} \left( \frac{(\text{number from molarity}) \text{ mol } 1}{10^3 \; \cancel{\text{mL solution } 1}} \right)$$

For a gas at STP, the molar volume at STP can be used as a conversion factor to convert from volume of gas at STP to moles.  **Objective 19**

$$\underline{\quad\quad} \; \cancel{\text{L } 1} \left( \frac{1 \text{ mol } 1}{22.414 \; \cancel{\text{L } 1}} \right)$$

For a gas under conditions other than STP, the ideal gas equation can be used to convert volume to moles, or the universal gas constant, $R$, can be used as a conversion factor.  **Objective 19**

$$n_1 = \frac{P_1 V_1}{R T_1}$$

or

$$\underline{\quad\quad} \; \cancel{\text{L } 1} \left( \frac{\cancel{K} \cdot \text{mol}}{0.082058 \; \cancel{\text{L} \cdot \text{atm}}} \right) \left( \frac{\underline{\quad} \; \cancel{\text{atm}}}{\underline{\quad} \; \cancel{K}} \right)$$

or

$$\underline{\quad\quad} \; \cancel{\text{L } 1} \left( \frac{\cancel{K} \cdot \text{mol}}{8.3145 \; \cancel{\text{L} \cdot \text{kPa}}} \right) \left( \frac{\underline{\quad} \; \cancel{\text{atm}}}{\underline{\quad} \; \cancel{K}} \right)$$

**Step 2**  Convert moles of substance 1 to moles of substance 2, using the coefficients from the balanced equation.

**Step 3**  Convert moles of substance 2 to amount of substance 2.

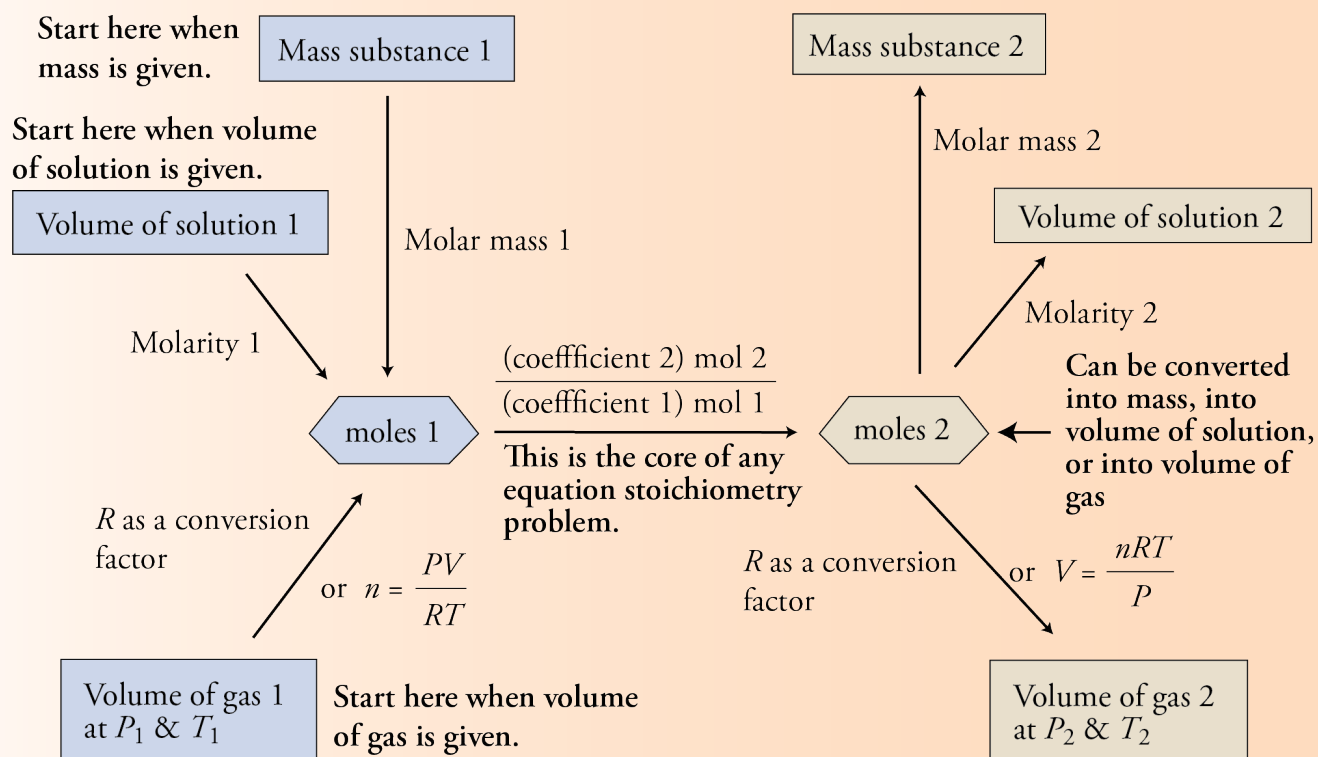For a pure solid or liquid, this will be a molar mass conversion from moles to mass.

$$\underline{\quad\quad} \left( \frac{(\text{coefficient 2}) \; \cancel{\text{mol } 2}}{(\text{coefficient 1}) \; \cancel{\text{mol } 1}} \right) \left( \frac{(\text{molar mass}) \text{ g } 2}{1 \; \cancel{\text{mol } 2}} \right)$$

**Sample Study Sheet 11.3**
Equation Stoichiometry Problems

For solutions, molarity can be used to convert from moles to volume of solution.

$$\underline{\quad} \left( \frac{\text{(coefficient 2)}\,\cancel{\text{mol 2}}}{\text{(coefficient 1)}\,\cancel{\text{mol 1}}} \right) \left( \frac{1\text{ L solution 2}}{\text{(number from molarity)}\,\cancel{\text{mol 2}}} \right)$$

$$\text{or}\quad \underline{\quad} \left( \frac{\text{(coefficient 2)}\,\cancel{\text{mol 2}}}{\text{(coefficient 1)}\,\cancel{\text{mol 1}}} \right) \left( \frac{10^3\text{ mL solution 2}}{\text{(number from molarity)}\,\cancel{\text{mol 2}}} \right)$$

**OBJECTIVE 19**

For a gas at STP, the molar volume at STP can be used as a conversion factor to convert from moles to volume of gas at STP.

$$\underline{\quad} \left( \frac{\text{(coefficient 2)}\,\cancel{\text{mol 2}}}{\text{(coefficient 1)}\,\cancel{\text{mol 1}}} \right) \left( \frac{22.414\text{ L 2}}{1\,\cancel{\text{mol 2}}} \right)$$

**OBJECTIVE 19**

For a gas under conditions other than STP, the ideal gas equation or the universal gas constant, $R$, can be used to convert from moles to volume of gas.

$$V_2 = \frac{n_2 R T_2}{P_2}$$

$$\text{or}\quad \underline{\quad} \left( \frac{\text{(coefficient 2)}\,\cancel{\text{mol 2}}}{\text{(coefficient 1)}\,\cancel{\text{mol 1}}} \right) \left( \frac{0.082058\text{ L}\cdot\cancel{\text{atm}}}{\cancel{\text{K}}\cdot\cancel{\text{mol}}} \right) \left( \frac{\underline{\quad}\,\cancel{\text{K}}}{\underline{\quad}\,\cancel{\text{atm}}} \right)$$

$$\text{or}\quad \underline{\quad} \left( \frac{\text{(coefficient 2)}\,\cancel{\text{mol 2}}}{\text{(coefficient 1)}\,\cancel{\text{mol 1}}} \right) \left( \frac{8.3145\text{ L}\cdot\cancel{\text{kPa}}}{\cancel{\text{K}}\cdot\cancel{\text{mol}}} \right) \left( \frac{\underline{\quad}\,\cancel{\text{K}}}{\underline{\quad}\,\cancel{\text{kPa}}} \right)$$

**EXAMPLES**   See Examples 11.5, 11.6, 11.7, and 11.8.

**Figure 11.10**
Summary of Equation Stoichiometry Possibilities



Start here when mass is given.
Start here when volume of solution is given.

Mass substance 1

Mass substance 2

Volume of solution 1

Molar mass 2

Volume of solution 2

Molarity 1

Molar mass 1

$$\frac{\text{(coeffficient 2) mol 2}}{\text{(coeffficient 1) mol 1}}$$

This is the core of any equation stoichiometry problem.

Molarity 2

Can be converted into mass, into volume of solution, or into volume of gas

moles 1

moles 2

$R$ as a conversion factor

or $n = \dfrac{PV}{RT}$

$R$ as a conversion factor

or $V = \dfrac{nRT}{P}$

Volume of gas 1 at $P_1$ & $T_1$

Start here when volume of gas is given.

Volume of gas 2 at $P_2$ & $T_2$

## EXERCISE 11.3 - Equation Stoichiometry

OBJECTIVE 19

Iron is combined with carbon in a series of reactions to form pig iron, which is about 4.3% carbon.

$$2C + O_2 \rightarrow 2CO$$
$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$
$$2CO \rightarrow C(in\ iron) + CO_2$$
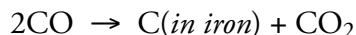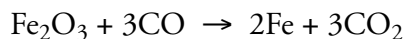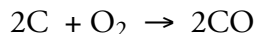
Pig iron is easier to shape than pure iron, and the presence of carbon lowers its melting point from the 1539 °C required to melt pure iron to 1130 °C.

a.   In the first reaction, what minimum volume of oxygen at STP is necessary to convert 125 Mg of carbon to carbon monoxide?

b.   In the first reaction, what is the maximum volume of carbon monoxide at 1.05 atm and 35 °C that could form from the conversion of $8.74 \times 10^5$ L of oxygen at 0.994 atm and 27 °C?

## EXERCISE 11.4 - Equation Stoichiometry

OBJECTIVE 19

Sodium hypochlorite, NaOCl, found in household bleaches, can be made from a reaction using chlorine gas and aqueous sodium hydroxide:

$$Cl_2(g) + 2NaOH(aq) \rightarrow NaOCl(aq) + NaCl(aq) + H_2O(l)$$
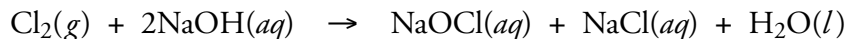
What minimum volume of chlorine gas at 101.4 kPa and 18.0 °C must be used to react with all the sodium hydroxide in 3525 L of 12.5 M NaOH?

## 11.4  Dalton's Law of Partial Pressures

Most gaseous systems contain a mixture of gases. Air is a mixture of nitrogen gas, oxygen gas, xenon gas, carbon dioxide gas, and many others. A typical "neon" light on a Las Vegas marquee contains argon gas as well as neon. The industrial reaction that forms nitric acid requires the mixing of ammonia gas and oxygen gas.

When working with a mixture of gases, we are sometimes interested in the total pressure exerted by all of the gases together, and sometimes we are interested in the portion of the total pressure that is exerted by only one of the gases in the mixture. The portion of the total pressure that one gas in a mixture of gases contributes is called the **partial pressure** of the gas. Distance runners find it harder to run at high altitudes because the partial pressure of oxygen in the air they breathe at sea level is about 21 kPa, but at 6000 ft above sea level, it drops to about 17 kPa. Table 11.1 shows the partial pressures of various gases in dry air. (The percentage of water vapor in air varies from place to place and day to day.)

**Table 11.1**

Gases in Dry Air at STP

| Gas | Percentage of total particles | Partial pressure (kPa) |
|---|---|---|
| nitrogen | 78.081 | 79.116 |
| oxygen | 20.948 | 21.226 |
| argon | 0.934 | 0.995 |
| carbon dioxide | 0.035 | 0.035 |
| neon | 0.002 | 0.002 |
| helium | 0.0005 | 0.0005 |
| methane | 0.0001 | 0.0001 |
| krypton | 0.0001 | 0.0001 |
| nitrous oxide | 0.00005 | 0.00005 |
| hydrogen | 0.00005 | 0.00005 |
| ozone | 0.000007 | 0.000007 |
| xenon | 0.000009 | 0.000009 |

OBJECTIVE 20

To get an understanding of the relationship between the total pressure of a mixture of gases and the partial pressure of each gas in the mixture, let's picture a luminous tube being filled with neon and argon gases. Let's say the neon gas is added to the tube first, and let's picture ourselves riding on one of the neon atoms. The particle is moving rapidly, colliding constantly with other particles and with the walls of the container. Each collision with a wall exerts a small force pushing out against the wall. The total pressure (force per unit area) due to the collisions of all of the neon atoms with the walls is determined by the rate of collision of the neon atoms with the walls and the average force per collision. The only ways that the neon gas pressure can be changed are either by a change in the rate of collisions with the walls or by a change in the average force per collision.

Now consider the effect on gas pressure of the addition of argon gas—at the same temperature—to the tube that already contains neon. Assuming that the mixture of gases acts as an ideal gas, there are no significant attractions or repulsions between the particles, and the volume occupied by the particles is very small. Except for collisions between particles, each gas particle acts independently of all of the other particles in the tube. Now that there are also a certain number of argon atoms in the tube, the neon atoms, including the one we are riding on, have more collisions with other particles and change their direction of motion more often, but they still collide with the walls at the same rate as before. If the temperature stays the same, the average velocity of the neon atoms and their average force per collision with the walls is the same.

If the neon atoms in the mixture are colliding with the walls of the tube at the same rate and with the same average force per collision as they did when alone, they are exerting the same pressure against the walls now as they did when alone. Therefore, the partial pressure of neon in the argon-neon mixture is the same as the pressure that the neon exerted when it was alone. Assuming ideal gas character, the **partial pressure** of any gas in a mixture is the pressure that the gas would exert if it were alone in the container.

Because the argon atoms in the mixture are also acting independently, the partial pressure of the argon gas is also the same as the pressure it would have if it were alone. The total pressure in the luminous tube that now contains both neon gas and argon gas is equal to the sum of their partial pressures (Figure 11.11).

$$P_{total} = P_{Ne} + P_{Ar}$$
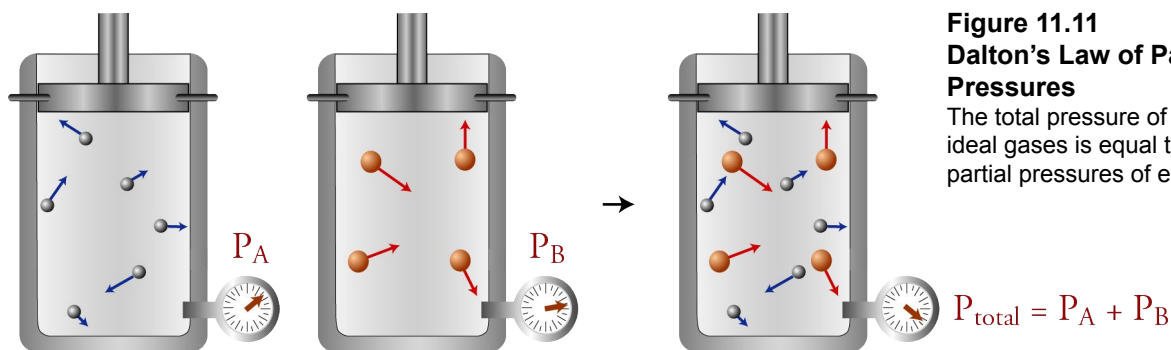
**Figure 11.11**
**Dalton's Law of Partial Pressures**
The total pressure of a mixture of ideal gases is equal to the sum of the partial pressures of each gas.

$$P_{total} = P_A + P_B$$

The situation is summarized in **Dalton's Law of Partial Pressures**: *the total pressure of a mixture of gases is equal to the sum of the partial pressures of all the gases.*

$$P_{total} = \Sigma P_{partial}$$

OBJECTIVE 21

Dalton's law of partial pressures can be combined with the ideal gas equation to solve many kinds of problems involving mixtures of gases. One way of combining them begins by rearranging the ideal gas equation to express the partial pressures of gases in a mixture—for example, the partial pressures of neon gas and argon gas in a mixture of argon and neon gases:

$$PV = nRT \quad \text{or} \quad P = \frac{nRT}{V}$$

Therefore,

$$P_{Ne} = \frac{n_{Ne}RT_{Ne}}{V_{Ne}} \qquad P_{Ar} = \frac{n_{Ar}RT_{Ar}}{V_{Ar}}$$

$$P_{total} = P_{Ne} + P_{Ar} = \frac{n_{Ne}RT_{Ne}}{V_{Ne}} + \frac{n_{Ar}RT_{Ar}}{V_{Ar}}$$

Because both neon and argon gases expand to fill the whole container, the volume for the argon gas and the volume for the neon gas are equal. Similarly, because the collisions between gas particles in the same container will lead the gases to have the same temperature, the temperatures for the argon gas and the neon gas are also equal. Therefore, the common $RT/V$ can be factored out to yield a simplified form of the equation.

$$V_{Ar} = V_{Ne} = V \quad T_{Ar} = T_{Ne} = T$$

$$P_{total} = \frac{n_{Ne}RT}{V} + \frac{n_{Ar}RT}{V} = (n_{Ne} + n_{Ar})\frac{RT}{V}$$

In other words, the total pressure of a mixture of gases is equal to the sum of the moles of all gases in the mixture times RT/V. The general form of this equation is

$$P_{total} = (\Sigma n_{each \, gas})\frac{RT}{V}$$

OBJECTIVE 22

The following sample study sheet describes calculations that can be done using Dalton's Law of Partial Pressures.

Sample Study Sheet 11.4
Using Dalton's Law of Partial Pressures

OBJECTIVE 21
OBJECTIVE 22

**TIP-OFF** – The problem involves a mixture of gases and no chemical reaction. You are asked to calculate a value for one of the variables in the equations below, and you are given (directly or indirectly) values for the other variables.

**GENERAL STEPS - THE FOLLOWING STEPS CAN BE USED TO WORK THESE PROBLEMS.**

**STEP 1**  Assign variables to the values that are given and the value that is unknown.

**STEP 2**  From the following equations, choose the one that best fits the variables assigned in Step 1.

$$P_{total} = \Sigma P_{partial} \quad \text{or} \quad P_{total} = (\Sigma n_{each\ gas})\frac{RT}{V}$$

**STEP 3**  Rearrange the equation to solve for your unknown.

**STEP 4**  Plug in the values for the given properties.

**STEP 5**  Make sure that the equation yields the correct units. Make any necessary unit conversions.

**STEP 6**  Calculate your answer and report it with the correct units and with the correct significant figures.

## EXAMPLE 11.9 - Dalton's Law of Partial Pressures

OBJECTIVE 21

A typical 100-watt light bulb contains a mixture of argon gas and nitrogen gas. In a light bulb with a total gas pressure of 111 kPa, enough argon is present to yield a partial pressure of 102 kPa. What is the partial pressure of the nitrogen gas?

*Solution*

**STEP 1**  Assign variables.

$$P_T = 111 \text{ kPa} \qquad P_{Ar} = 102 \text{ kPa} \qquad P_{N_2} = ?$$

**STEP 2**  Our variables fit the general equation

$$P_{total} = \Sigma P_{partial} \quad \text{or} \quad P_{total} = P_{Ar} + P_{N_2}$$

**STEPS 3-6**  Solve the equation for the unknown variable, plug in the given values, check the units, and calculate the answer.

$$P_{N_2} = P_{total} - P_{Ar} = 111 \text{ kPa} - 102 \text{ kPa} = \textbf{9 kPa}$$

## EXAMPLE 11.10 - Dalton's Law of Partial Pressures

If an 85-mL light bulb contains 0.140 grams of argon and 0.011 grams of nitrogen at 20 °C, what is the total pressure of the mixture of gases?

OBJECTIVE 22

*Solution*

**STEP 1**  Although there are no mass variables in our partial pressure equations, we do know that we can convert mass to moles using molar masses.

$$n_{Ar} = ? \text{ mol Ar} = 0.140 \text{ g Ar} \left( \frac{1 \text{ mol Ar}}{39.948 \text{ g Ar}} \right) = 0.00350 \text{ mol Ar}$$

$$n_{N_2} = ? \text{ mol N}_2 = 0.011 \text{ g N}_2 \left( \frac{1 \text{ mol N}_2}{28.0134 \text{ g N}_2} \right) = 0.00039 \text{ mol N}_2$$

$$V = 85 \text{ mL} \qquad T = 20 °C + 273.15 = 293 \text{ K}$$

**STEP 2**  Our variables fit the following form of Dalton's Law of Partial Pressures.

$$P_{total} = (\Sigma n_{each \, gas}) \frac{RT}{V} \qquad P_{total} = (n_{Ar} + n_{N_2}) \frac{RT}{V}$$

**STEPS 3-6**  We do not need to rearrange the equation algebraically, but when we plug in our values with their units, we see that we need to convert 85 mL to liters to get our units to cancel correctly.

$$P_{total} = (0.00350 \text{ mol Ar} + 0.00039 \text{ mol N}_2) \left( \frac{\left( \frac{0.082058 \text{ L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \right) (293 \text{ K})}{85 \text{ mL}} \right) \left( \frac{10^3 \text{ mL}}{1 \text{ L}} \right)$$

$$= 1.1 \text{ atm (or } 1.1 \times 10^2 \text{ kPa)}$$

## EXERCISE 11.5 - Dalton's Law of Partial Pressures

A typical "neon light" contains neon gas mixed with argon gas.

OBJECTIVE 21
OBJECTIVE 22

   a.  If the total pressure of the mixture of gases is 1.30 kPa and the partial pressure of neon gas is 0.27 kPa, what is the partial pressure of the argon gas?
   b.  If 6.3 mg of Ar and 1.2 mg Ne are added to a 375-mL tube at 291 K, what is the total pressure of the gases in millimeters of mercury?

# Special Topic 11.1      A Greener Way to Spray Paint

United States industries use an estimated 1.5 billion liters of paints and other coatings per year, and much of this is applied by spraying. Each liter sprayed from a canister releases an average of 550 grams of volatile organic compounds (VOCs). Some of these VOCs are hazardous air pollutants.

The mixture that comes out of the spray can has two kinds of components: (1) the solids being deposited on the surface as a coating and (2) a solvent blend that allows the solids to be sprayed and to spread evenly. The solvent blend must dissolve the coatings into a mixture that has a thin enough consistency to be easily sprayed. But a mixture that is thin enough to be easily sprayed will be too runny to remain in place when deposited on a surface. Therefore, the solvent blend contains additional components that are so volatile they will evaporate from the spray droplets between the time the spray leaves the spray nozzle and the time the spray hits the surface. Still other, slower-evaporating components do not evaporate until after the spray hits the surface. They remain in the coating mixture long enough to cause it to spread out evenly on the surface. Since the more volatile solvents have escaped, the mixture that hits the surface is thick enough not to run or sag.

The Clean Air Act has set strict limitations on the emission of certain VOCs, so new safer solvents are needed to replace them. One new spray system has been developed that yields a high-quality coating while emitting as much as 80% fewer VOCs of all types and none of the VOCs that are considered hazardous air pollutants. This system is called the supercritical fluid spray process. The solvent mixture for this process still contains some of the slow evaporating solvents that allow the coating to spread evenly but replaces the fast evaporating solvents with high pressure $CO_2$.

Some gases can be converted to liquids at room temperature by being compressed into a smaller volume, but for each gas, there is a temperature above which the particles are moving too fast to allow a liquid to form no matter how much the gas is compressed. The temperature above which a gas cannot be liquefied is called the critical temperature. The very high pressures that are possible for a gas above its critical temperature allow the formation of gas systems in which the density approaches the densities of liquids. At this high pressure and density, the substance takes on characteristics of both gases and liquids and is called a supercritical fluid.

The critical temperature of $CO_2$ is 31 °C. Above this temperature, carbon dioxide can be compressed to a very high pressure and relatively high density supercritical fluid. Like a liquid, the supercritical carbon dioxide will mix with or dissolve the blend of coating and low volatility solvent to form a product that is thin enough to be sprayed easily, in very small droplets. The supercritical $CO_2$ has a very high volatility, so it evaporates from the droplets almost immediately after they leave the spray nozzle, leaving a mixture that is thick enough not to run or sag when it hits the surface. The mixture is sprayed at temperatures of about 50 °C and a pressure of 100 atm (about 100 times normal room pressure).

Because carbon dioxide is much less toxic than the VOCs it replaces and because it is nonflammable and relatively inert, it is much safer to use in the workplace. It is also far less expensive. Moreover, because the $CO_2$ can be obtained from the production of other chemicals, the new process does not lead to an increase in carbon dioxide in the atmosphere. In fact, since the VOCs that it replaces can form $CO_2$ in the atmosphere, the supercritical fluid spray process actually leads to a decrease in atmospheric $CO_2$ levels compared to what they would be with the traditional spray process.

Source: M. D. Donohue, J. L. Geiger, A. A. Kiamos, and K. A. Nielsen, "Reduction of VOC Emissions During Spray Painting." In *Green Chemistry*; ed. P. T. Anastas and T. C. Williams (American Chemical Society: Washington DC, 1996), pp. 153-167.



By taking advantage of the properties of gases at high temperatures and pressures, scientists have invented a new, environmentally friendly spray-painting.

# SPECIAL TOPIC 11.2    Green Decaf Coffee

Caffeine, the primary stimulant found naturally in coffee, tea, and chocolate, is artificially added to many other products, such as soft drinks, over the counter stimulants, cold remedies, and pain relievers. It is the world's most widely used drug. Different types of coffee contain different amounts of caffeine, and different brewing techniques lead to further variations in content. Depending on the type of coffee (an arabica coffee contains about half the caffeine of a robusta) and the fineness of the grounds, the drip method of brewing leads to 110 to 150 mg of caffeine per cup. Instant coffee has 40-108 mg of caffeine per cup. The average cup of coffee contains from 40 to 150 mg of caffeine.

Some people like the taste of coffee and the rituals associated with it (the customary cup of coffee after dinner) but do not like the effects of the caffeine. They may choose to drink decaffeinated coffee, which has 2-4 mg of caffeine per cup. A coffee must have at least 97% of its caffeine removed to be considered decaffeinated coffee.

One of the commercial methods for decaffeinating coffee is the direct-contact method. The unroasted coffee beans are first softened with steam and then brought in direct contact with a decaffeinating agent, such as dichloromethane (most often called methylene chloride in this context), $CH_2Cl_2$. The caffeine dissolves in the dichloromethane, after which the dichloromethane/caffeine solution is removed from the beans.

There is some evidence that dichloromethane is carcinogenic. According to the United States Food and Drug administration (FDA), most decaffeinated coffee has less than 0.1 parts per million (ppm) of residual dichloromethane, which is 100 times less than the maximum level of 10 ppm allowed by the FDA. Despite this low residual concentration of dichloromethane, there is still some concern about a possible health hazard, so substitutes are being sought for many of its applications.

One alternative to the current methods of decaffeination uses supercritical carbon dioxide. The unroasted beans are softened by steam and then immersed in carbon dioxide at very high temperature and pressure. It penetrates the beans more than a typical liquid would, and then, like a liquid, dissolves the caffeine. When the supercritical carbon dioxide is later separated from the beans, it carries about 97% of the caffeine along with it. Any residue of carbon dioxide remaining on the beans is quickly lost as gaseous $CO_2$. The workers employed in this method of decaffeination are safer without the contact with dichloromethane, and the coffee drinkers do not have to worry about the very slight residue of dichloromethane that remains when the coffee is decaffeinated using the direct-contact method.

All of these products contain caffeine.

## Chapter Glossary

**Ideal gas model**   The model for gases that assumes (1) the particles are point-masses (they have mass but no volume) and (2) there are no attractive or repulsive forces between the particles.

**Ideal gas**   A gas for which the ideal gas model is a good description.

**Pressure**   Force per unit area.

**Boyle's Law**   The pressure of a gas is inversely proportional to the volume it occupies if the number of gas particles and the temperature are constant.

**Gay-Lussac's Law**   The pressure of an ideal gas is directly proportional to the Kelvin temperature of the gas if the volume and the number of gas particles are constant.

**Charles' Law**   Volume and temperature are directly proportional if the number of gas particles and pressure are constant.

**Avogadro's Law**   Volume and the number of gas particles are directly proportional if the temperature and pressure are constant.

**Universal gas constant, $R$**   The constant in the ideal gas equation.

**Partial pressure**   The portion of the total pressure that one gas in a mixture of gases contributes. Assuming ideal gas character, the partial pressure of any gas in a mixture is the pressure that the gas would yield if it were alone in the container.

**Dalton's Law of Partial Pressures**   The total pressure of a mixture of gases is equal to the sum of the partial pressures of each gas.

**You can test yourself on the glossary terms at the textbook's Web site.**

## Chapter Objectives

**The goal of this chapter is to teach you to do the following.**

1. Define all of the terms in the Chapter Glossary.

**Section 11.1 Gases and Their Properties**

2. For a typical gas, state the percentage of space inside a gas-filled container that is occupied by the gas particles themselves.
3. State the average distance traveled between collisions by oxygen molecules, $O_2$, at normal room temperature and pressure.
4. Write the key assumptions that distinguish an ideal gas from a real gas.
5. Describe the process that gives rise to gas pressure.
6. State the accepted SI unit for gas pressure.
7. Convert between the names and abbreviations for the following pressure units: pascal (Pa), atmosphere (atm), millimeter of mercury (mmHg), and torr.
8. Convert a gas pressure described in pascals (Pa), atmospheres (atm), millimeters of mercury (mmHg), or torr to any of the other units.
9. Convert between the names and variables used to describe pressure ($P$), temperature, ($T$), volume ($V$), and moles of gas ($n$).
10. For each of the following pairs of gas properties, describe the relationship between the properties, describe a simple system that could be used to demonstrate the relationship, and explain the reason for the relationship: (a) volume and pressure when the number of gas particles and temperature are constant, (b) pressure and temperature when volume and the number of gas particles are constant, (c) volume and temperature when pressure and the number of gas particles are constant, (d) the number of gas particles and pressure when volume and temperature are constant, and (e) the number of gas particles and volume when pressure and temperature are constant.

11. Explain why decreased volume for the gasoline-air mixture in the cylinders of a gasoline engine, increased number of gas particles, and increased temperature lead to an increase in pressure in the cylinders.
12. Explain why air moves in and out of our lungs as we breathe.
13. Explain why balloons expand as they rise into the atmosphere.

### Section 11.2  Ideal Gas Calculations

14. Write at least one value for the universal gas constant, $R$, including units.
15. Given values for three of the following four variables, calculate the fourth: $P$, $V$, $n$, and $T$.
16. Given values for four out of the following five variables, calculate the fifth: $P$, $V$, $g$, $M$, and $T$.
17. Given the pressure and temperature of a gas, calculate its density.
18. Given (directly or indirectly) values for seven out of the following eight variables, calculate the eighth: $P_1$, $V_1$, $n_1$, $T_1$, $P_2$, $V_2$, $n_2$, and $T_2$.

### Section 11.3  Equation Stoichiometry and Ideal Gases

19. Convert between volume of gas and moles of gas in an equation stoichiometry problem using either the molar volume at STP, the ideal gas equation, or $R$ as a conversion factor.

### Section 11.4  Dalton's Law of Partial Pressures

20. Explain why the total pressure of a mixture of ideal gases is equal to the sum of the partial pressures of each gas.
21. Given all but one of the following properties for a mixture of gases, calculate the one not given: total pressure of the mixture of gases and partial pressure of each gas.
22. Given all but one of the following properties for a mixture of gases, calculate the one not given: total pressure of the mixture of gases, mass or moles of each gas, temperature, and volume of the container.

---

**1.** Describe the particle nature of gases.
**2.** What is 265.2 °C on the Kelvin scale? Convert 565.7 K into °C.
**3.** About 55% of industrially produced sodium sulfate is used to make detergents. It is made from the following reaction.

$$4NaCl + 2SO_2 + 2H_2O + O_2 \;\rightarrow\; 2Na_2SO_4 + 4HCl$$
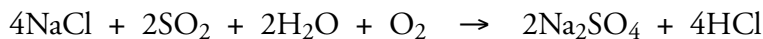
   a. What is the maximum mass of sodium sulfate that can be produced in the reaction of 745 Mg of sodium chloride with excess $SO_2$, $H_2O$, and $O_2$?
   b. What is the maximum mass of sodium sulfate that can be produced in the reaction of 745 Mg of sodium chloride with 150 Mg $H_2O$ and excess $SO_2$ and $O_2$?
   c. If 868 Mg $Na_2SO_4$ are formed in the reaction of 745 Mg of sodium chloride with 150 Mg $H_2O$ and excess $SO_2$ and $O_2$, what is the percent yield?
   d. What volume of 2.0 M $Na_2SO_4$ can be formed from 745 Mg of sodium chloride with excess $SO_2$, $H_2O$, and $O_2$?

**Review Questions**

# Key Ideas

Complete the following statements by writing one of these words or phrases in each blank.

| | |
|---|---|
| 0 °C | increases |
| $10^{-7}$ m | Kelvin |
| 0.1% | Kelvin temperature |
| 70% | liters, L |
| 99.9% | molar mass |
| 1 atm | molarity |
| 101.325 kPa | moles, mol |
| 273.15 K | moles of gas |
| 760 mmHg | necessary unit conversions |
| algebra | number of collisions |
| alone | number of particles |
| attractions | pascal, Pa |
| attractive or repulsive | point-masses |
| Avogadro's Law | portion |
| collide | pressure |
| continuous | rapid |
| correct equation | real |
| degrees Celsius, °C | repulsions |
| direction of motion | sum |
| directly | temperature |
| empty space | temperature and volume |
| force | ten |
| force per unit area | velocity |
| greater force | volume |

4. Under typical conditions, the average distance between gas particles is about _____ times their diameter.

5. Because of the large distances between gas particles, the volume occupied by the particles themselves is very small compared to the volume of the _____ around them.

6. For a gas at room temperature and pressure, the gas particles themselves occupy about _____ of the total volume. The other _____ of the total volume is empty space (whereas in liquids, about _____ of the volume is occupied by particles).

7. Because of the large distances between gas particles, there are very few _____ or _____ between them.

8. The particles in a gas are in _____ and _____ motion.

9. As the temperature of a gas increases, the particles' velocity _____.

10. The particles in a gas are constantly colliding with the walls of the container and with each other. Because of these collisions, the gas particles are constantly changing their _____ and their _____.

11. Oxygen, $O_2$, molecules at normal temperatures and pressures move an average of _____ between collisions.

12. The particles of an ideal gas are assumed to be _____, that is, particles that have a mass but occupy no volume.

13. There are no _____ forces at all between the particles of an ideal gas.

**14.** The ideal gas model is used to predict changes in four related gas properties: _____, _____, _____, and _____.

15. Volumes of gases are usually described in _____ or cubic meters, $m^3$, and numbers of particles are usually described in _____.

**16.** Although gas temperatures are often measured with thermometers that report temperatures in _____ scientists generally use _____ temperatures for calculations.

17. Each time a gas particle collides with and ricochets off one of the walls of its container, it exerts a(n) _____ against the wall. The sum of the forces of these ongoing collisions of gas particles against all the container's interior walls creates a continuous pressure upon those walls.

**18.** The accepted SI unit for gas pressure is the _____.

19. _____ gases deviate somewhat from predicted behavior of ideal gases.

**20.** The observation that the pressure of an ideal gas is inversely proportional to the volume it occupies if the _____ and the temperature are constant is a statement of Boyle's Law.

21. When the volume of a chamber that contains a gas decreases but the number of gas particles remains constant, there is an increase in the concentration (number of gas particles per liter) of the gas. This leads to an increase in the number of particles near any given area of the container walls at any time and to an increase in the number of collisions against the walls per unit area in a given time. More collisions mean an increase in the _____, or pressure, of the gas.

**22.** The pressure of an ideal gas is directly proportional to the _____ of the gas if the volume and the number of gas particles are constant. This relationship is sometimes called Gay-Lussac's Law.

23. Increased temperature means increased motion of the particles of a gas. If the particles are moving faster in the container, they will _____ with the walls more often and with _____ per collision. This leads to a greater overall force pushing on the walls and to a greater force per unit area or pressure.

**24.** For an ideal gas, volume and temperature described in kelvins are _____ proportional if the number of gas particles and pressure are constant. This is a statement of Charles' Law.

25. The increase in the number of gas particles in a constant-volume container leads to an increase in the _____ with the walls per unit time. This leads to an increase in the force per unit area–that is, to an increase in gas pressure.

**26.** If the _____ of an ideal gas are held constant, the number of particles in a container and the gas pressure are directly proportional.

27. The relationship between the number of particles of an ideal gas and volume is summarized by _____, which states that the volume and the number of gas particles are directly proportional if the temperature and pressure are constant.

**28.** It is always a good idea to include the units in a solved equation as well as the numbers. If the units cancel to yield a reasonable unit for the unknown property, you can feel confident that you picked the _____, that you did the _____ correctly to solve for your unknown, and that you have made the _____.

29. STP stands for "standard temperature and pressure." Standard temperature is
_____, or _____, and standard pressure is _____,
or _____, or _____.

**30.** There are three different ways to convert between a measurable property and
moles in equation stoichiometry problems. For pure liquids and solids, we can
convert between mass and moles, using the _____ as a conversion
factor. For solutions, _____ provides a conversion factor that enables
us to convert between moles of solute and volume of solution.

31. The _____ of the total pressure that one gas in a mixture of gases
contributes is called the partial pressure of the gas.

**32.** Assuming ideal gas character, the partial pressure of any gas in a mixture is the
pressure that the gas would exert if it were _____ in the container.

33. Dalton's law of partial pressures states that the total pressure of a mixture of gases
is equal to the _____ of the partial pressures of gas all the gases.

## Chapter Problems

### Section 11.1 Gases and their Properties

OBJECTIVE 2

**34.** For a gas under typical conditions, what approximate percentage of the volume is
occupied by the gas particles themselves?

35. One of the car maintenance tasks you might do at a service station is to fill your
tires with air. Why are the tires not really full when you get to the desired tire
pressure? Why can you still add more air to the tire?

**36.** Why is it harder to walk through water than to walk through air?

OBJECTIVE 3

37. What is the average distance traveled between collisions by molecules of oxygen
gas, $O_2$, at room temperature and pressure?

OBJECTIVE 4
OBJECTIVE 5

**38.** What are the key assumptions that distinguish an ideal gas from a real gas?

39. Consider a child's helium-filled balloon. Describe the process that creates the gas
pressure that keeps the balloon inflated.

**40.** A TV weather person predicts a storm for the next day and refers to the dropping
barometric pressure to support the prediction. What causes the pressure in the
air?

OBJECTIVES 6, & 7

41. What is the accepted SI unit for gas pressure? What is its abbreviation?

OBJECTIVE 8

**42.** The pressure at the center of the earth is $4 \times 10^{11}$ pascals, Pa. What is this
pressure in kPa, atm, mmHg, and torr?

OBJECTIVE 8

43. The highest sustained pressure achieved in the laboratory is $1.5 \times 10^7$ kilopascals,
kPa. What is this pressure in atm, mmHg, and torr?

**44.** What does inversely proportional mean?

OBJECTIVE 10

45. For each of the following pairs of gas properties, describe the relationship
between the properties, describe a simple system that could be used to
demonstrate the relationship, and explain the reason for the relationship:
(a) volume and pressure when number of gas particles and temperature are
constant, (b) pressure and temperature when volume and the number of gas
particles are constant, (c) volume and temperature when pressure and the number
of gas particles are constant, (d) the number of gas particles and pressure when
volume and temperature are constant, and (e) the number of gas particles and
volume when pressure and temperature are constant.

46. Explain why decreased volume for the gasoline-air mixture in the cylinders of a gasoline engine, increased number of gas particles, and increased temperature lead to an increase in pressure in the cylinders.

OBJECTIVE 11

47. Ammonia, $NH_3$, is a gas that can be dissolved in water and used as a cleaner. When an ammonia solution is used to clean the wax off a floor, ammonia gas escapes from the solution, mixes easily with the gas particles in the air, and spreads throughout the room. Very quickly, everyone in the room can smell the ammonia gas. Explain why gaseous ammonia mixes so easily with air.

48. After gasoline vapor and air are mixed in the cylinders of a car, the piston in the cylinder moves up into the cylinder and compresses the gaseous mixture. Explain why these gases can be compressed.

49. Explain why air moves in and out of our lungs as we breathe.

OBJECTIVE 12

50. Explain why balloons expand as they rise into the atmosphere.

OBJECTIVE 13

51. With reference to the relationships between properties of gases, answer the following questions.

    a. In a common classroom demonstration, water is heated in a one gallon can, covered with a tight lid, and allowed to cool. As the can cools, it collapses. Why?

    b. Dents in ping pong balls can often be removed by placing the balls in hot water. Why?

52. Use the relationships between properties of gases to answer the following questions.

    a. When a balloon is placed in the freezer, it shrinks. Why?

    b. Aerosol cans often explode in fires. Why?

    c. Why do your ears pop when you drive up a mountain?

## Section 11.2  Ideal Gas Calculations

53. Neon gas produced for use in luminous tubes comes packaged in 1.0-L containers at a pressure of 1.00 atm and a temperature of 19 °C. How many moles of Ne do these cylinders hold?

OBJECTIVE 15

54. Fluorescent light bulbs contain a mixture of mercury vapor and a noble gas such as argon. How many moles of argon are in a 121-mL fluorescent bulb with a pressure of argon gas of 325 Pa at 40 °C?

OBJECTIVE 15

55. A typical aerosol can is able to withstand 10-12 atm without exploding.

    a. If a 375-mL aerosol can contains 0.062 moles of gas, at what temperature would the gas pressure reach 12 atm of pressure?

    b. Aerosol cans usually contain liquids as well as gas. If the 375-mL can described in part (a) contained liquid along with the 0.062 moles of gas, why would it explode at a lower temperature than if it contained only the gas? (Hint: What happens to a liquid when it is heated?)

OBJECTIVE 15

56. Ethylene oxide is produced from the reaction of ethylene and oxygen at 270-290 °C and 8-20 atm. In order to prevent potentially dangerous pressure buildups, the container in which this reaction takes place has a safety valve set to release gas when the pressure reaches 25 atm. If a 15-$m^3$ reaction vessel contains $7.8 \times 10^3$ moles of gas, at what temperature will the pressure reach 25 atm? (There are $10^3$ L per $m^3$.)
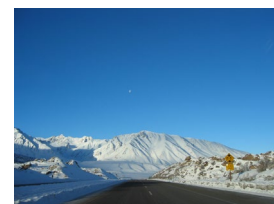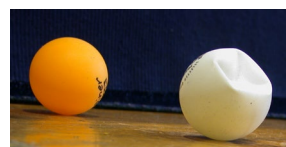
OBJECTIVE 15

OBJECTIVE 15



**57.** Bromomethane, $CH_3Br$, commonly called methyl bromide, is used as a soil fumigant, but because it is a threat to the ozone layer, its use is being phased out. This colorless gas is toxic by ingestion, inhalation, or skin absorption, so the American Conference of Government Industrial Hygienists has assigned it a threshold limit value, or TLV, a maximum concentration to which a worker may be repeatedly exposed day after day without experiencing adverse effects. (TLV standards are meant to serve as guides in control of health hazards, not to provide definitive dividing lines between safe and dangerous concentrations.) Bromomethane reaches its TLV when 0.028 moles escape into a room that has a volume of 180 m³ and a temperature of 18 °C. What is the pressure in atmospheres of the gas under those conditions? (There are $10^3$ L per m³.)

OBJECTIVE 15

58. Sulfur hexafluoride is a very stable substance that is used as a gaseous insulator for electrical equipment. Although $SF_6$ is considered relatively safe, it reaches its TLV (defined in question 57) when 143 moles of it escape into a room with a volume of 3500 m³ and a temperature of 21 °C. What is its pressure in pascals for these conditions?

OBJECTIVE 15

**59.** Gaseous chlorine, $Cl_2$, used in water purification, is a dense, greenish-yellow substance with an irritating odor. It is formed from the electrolysis of molten sodium chloride. Chlorine reaches its TLV (defined in question 57) in a laboratory when 0.017 moles of $Cl_2$ are released yielding a pressure of $7.5 \times 10^{-4}$ mmHg at a temperature of 19 °C. What is the volume in liters of the room?

OBJECTIVE 15

60. Hydrogen bromide gas, HBr, is a colorless gas used as a pharmaceutical intermediate. Hydrogen bromide reaches its TLV (defined in question 57) in a laboratory room when 0.33 moles of it are released yielding a pressure of 0.30 Pa at a temperature of 22 °C. What is the volume in m³ of the room?

OBJECTIVE 16

**61.** Hydrogen sulfide, $H_2S$, used to precipitate sulfides of metals, is a colorless gas that smells like rotten eggs. It can be detected by the nose at about 0.6 mg/m³. To reach this level in a 295 m³ room, 0.177 g $H_2S$ must escape into the room. What is the pressure in pascals of 0.177 g $H_2S$ that escapes into a 295-m³ laboratory room at 291 K?

OBJECTIVE 16

62. Nitrogen dioxide, $NO_2$, is a reddish-brown gas above 21.1 °C and a brown liquid below 21.1 °C. It is used to make nitric acid and as an oxidizer for rocket fuels. Its TLV (defined in question 57) is reached when 1.35 g of $NO_2$ escape into a room that has a volume of 225 m³. What is the pressure in millimeters of mercury of 1.35 g of $NO_2$ at 302 K in a 225-m³ room?

OBJECTIVE 16

**63.** Hydrogen chloride, HCl, is used to make vinyl chloride, which is then used to make polyvinyl chloride (PVC) plastic. Hydrogen chloride can be purchased in liquefied form in cylinders that contain 8 lb HCl. What volume in liters will 8.0 lb HCl occupy if the compound becomes a gas at 1.0 atm and 85 °C?

64. Sulfur dioxide, $SO_2$, forms in the combustion of the sulfur found in fossil fuels such as coal. In the air, $SO_2$ forms $SO_3$, which dissolves in water to form sulfuric acid. Thus $SO_2$ is a major contributor to acid rain as well as a strong eye and lung irritant. The 1990 amendments to the Clean Air Act of 1967 call for a reduction in sulfur dioxide released from power plants to 10 million tons per year. This is about one-half the emission measured in 1990. The US atmospheric standard for $SO_2$ is 0.35 mg/m$^3$, which would yield a pressure due to sulfur dioxide of 0.013 Pa at 17 °C. To what volume in m$^3$ must $1.0 \times 10^7$ tons of $SO_2$ expand in order to yield a pressure of 0.013 Pa at 17 °C? (There are 2000 lb per ton.)

65. The bulbs used for fluorescent lights have a mercury gas pressure of 1.07 Pa at 40 °C. How many milligrams of liquid mercury must evaporate at 40 °C to yield this pressure in a 1.39-L fluorescent bulb?

66. Low-pressure sodium lamps have a sodium vapor pressure of 0.7 Pa at 260 °C. How many micrograms of Na must evaporate at 260 °C to yield a pressure of 0.7 Pa in a 27-mL bulb?

67. Flashtubes are light bulbs that produce a high-intensity, very short duration flash of light. They are used in photography and for airport approach lighting. The xenon gas, Xe, in flashtubes has large atoms that can quickly carry the heat away from the filament to stop the radiation of light. At what temperature does 1.80 g Xe in a 0.625-L flashtube have a pressure of 75.0 kPa?

68. As part of the production of vinyl chloride—which is used to make polyvinyl chloride (PVC) plastic—1,2-dichloroethane, $ClCH_2CH_2Cl$, is decomposed to form vinyl chloride and hydrogen chloride. If 515 kg of 1,2-dichloroethane gas are released into a 96 m$^3$ container, at what temperature will the pressure become 3.4 atm?

69. When 0.690 g of an unknown gas are held in an otherwise-empty 285-mL container, the pressure is 756.8 mmHg at 19 °C. What is the molecular mass of the gas?

70. A 1.02 g sample of an unknown gas in a 537-mL container generates a pressure of 101.7 kPa at 298 K. What is the molecular mass of the gas?

71. Consider two helium tanks used to fill children's balloons. The tanks have equal volume and are at the same temperature, but tank A is new and therefore has a higher pressure of helium than tank B, which has been used to fill many balloons. Which tank holds gas with a greater density, tank A or tank B? Support your answer.

72. Two hot-air balloons have just been launched. They have equal volume and pressure, but balloon A has a more efficient heating system, so the gas in balloon A is at a higher temperature than the gas in balloon B. Which balloon holds gas with a greater density, balloon A or balloon B? Support your answer. Which balloon do you think will rise faster?

OBJECTIVE 17

**73.** Butadiene, $CH_2CHCHCH_2$, a suspected carcinogen, is used to make plastics such as styrene-butadiene rubber (SBR). Because it is highly flammable and forms explosive peroxides when exposed to air, it is often shipped in insulated containers at about 2 °C. What is the density of butadiene when its pressure is 102 kPa and its temperature is 2 °C?

74. Picture a gas in the apparatus shown in Figure 11.2. If the temperature and the number of gas particles remain constant, how would you change the volume to increase the pressure of the gas?

OBJECTIVE 18

**75.** In order to draw air into your lungs, your diaphragm and other muscles contract, increasing the lung volume. This lowers the air pressure of the lungs to below atmospheric pressure, and air flows in. When your diaphragm and other muscles relax, the volume of the lungs decreases, and air is forced out.

    a. If the total volume of your lungs at rest is 6.00 L and the initial pressure is 759 mmHg, what is the new pressure if the lung volume is increased to 6.02 L?

    b. If the total volume of your lungs at rest is 6.00 L and the initial pressure is 759 mmHg, at what volume will the pressure be 763 mmHg?

OBJECTIVE 18

76. Dry air is 78% nitrogen gas, 21% oxygen gas, and 0.9% argon gas. These gases can be separated by distillation after the air is first converted to a liquid. In one of the steps in this process, filtered air is compressed enough to increase the pressure to about 5.2 atm.

    a. To what volume must 175 L of air at 1.0 atm be compressed to yield a pressure of 5.2 atm?

    b. The argon derived from the distillation of air can be used in fluorescent light bulbs. When 225 L of argon gas at 0.997 atm are distilled from the air and allowed to expand to a volume of 91.5 m$^3$, what is the new pressure of the gas in pascals?

OBJECTIVE 18

**77.** A scuba diver has 4.5 L of air in his lungs 66 ft below the ocean surface, where the pressure is 3.0 atm. What would the volume of this gas be at the surface, where the pressure is 1.0 atm? If the diver's lungs can hold no more than 7.0 L without rupturing, is he in trouble if he doesn't exhale as he rises to the surface?

**78.** Picture a gas in the apparatus shown in Figure 11.2. If the volume and temperature remain constant, how would you change the number of gas particles to increase the pressure of the gas?

79. Picture a gas in the apparatus shown in Figure 11.2. If the pressure and temperature remain constant, how would you change the number of gas particles to increase the volume of the gas?

OBJECTIVE 18

**80.** A balloon containing 0.62 moles of gas has a volume of 15 L. Assuming constant temperature and pressure, how many moles of gas does the balloon contain when enough gas leaks out to decrease the volume to 11 L?

81. Picture a gas in the apparatus shown in Figure 11.2. If the volume and the number of gas particles remain constant, how would the temperature have to change to increase the pressure of the gas?

**82.** Consider a weather balloon with a volume on the ground of 113 m$^3$ at 101 kPa and 14 °C.

    a.  If the balloon rises to a height where the temperature is –50 °C and the pressure is 2.35 kPa, what will its volume be? (In 1958, a balloon reached 101,500 ft or 30.85 km, the height at which the temperature is about –50 °C.)

    b.  When the balloon rises to 10 km, where the temperature is –40 °C, its volume increases to 129 m$^3$. If the pressures inside and outside the balloon are equal, what is the atmospheric pressure at this height?

    c.  If the balloon has a volume of 206 m$^3$ at 25 km, where the pressure is 45 kPa, what is the temperature at this height?

**83.** Ethylene oxide, which is used as a rocket propellant, is formed from the reaction of ethylene gas with oxygen gas.

    a.  After 273 m$^3$ of ethylene oxide are formed at 748 kPa and 525 K, the gas is cooled at constant volume to 293 K. What will the new pressure be?

    b.  After 273 m$^3$ of ethylene oxide at 748 kPa and 525 K is cooled to 293 K, it is allowed to expand to $1.10 \times 10^3$ m$^3$. What is the new pressure?

    c.  What volume of ethylene gas at 293 K and 102 kPa must be compressed to yield 295 m$^3$ of ethylene gas at 808 kPa and 585 K?

**84.** The hydrogen gas used to make ammonia can be made from small hydrocarbons such as methane in the so-called steam-reforming process, run at high temperature and pressure.

    a.  If 3.2 atm of methane at 21 °C are introduced into a container, to what temperature must the gas be heated to increase the pressure to 12 atm?

    b.  If $4.0 \times 10^3$ L of methane gas at 21 °C is heated and allowed to expand at a constant pressure, what will the volume become in m$^3$ when the temperature reaches 815 °C?

    c.  What volume of methane gas at 21 °C and 1.1 atm must be compressed to yield $1.5 \times 10^4$ L of methane gas at 12 atm and 815 °C?

**Section 11.3  Equation Stoichiometry and Ideal Gases**

**85.** Chlorine gas is used in the production of many other chemicals, including carbon tetrachloride, polyvinyl chloride plastic, and hydrochloric acid. It is produced from the electrolysis of molten sodium chloride. What minimum mass of sodium chloride in megagrams is necessary to make $2.7 \times 10^5$ L of chlorine gas at standard temperature and pressure, STP?

$$2NaCl(l) \xrightarrow{\text{Electrolysis}} 2Na(l) + Cl_2(g)$$
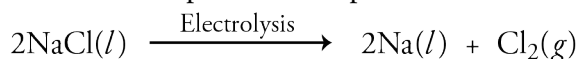
**86.** Sulfur is used in the production of vulcanized rubber, detergents, dyes, pharmaceuticals, insecticides, and many other substances. The final step in the Claus method for making sulfur is

$$SO_2(g) + 2H_2S(g) \rightarrow 3S(s) + 2H_2O(l)$$
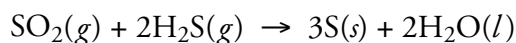
    a.  What minimum volume of sulfur dioxide at STP would be necessary to produce 82.5 kg of sulfur?

    b.  What maximum mass of sulfur in kilograms could be produced by the reaction of $1.3 \times 10^4$ m$^3$ of $SO_2(g)$ with $2.5 \times 10^4$ m$^3$ of $H_2S(g)$, both at STP?
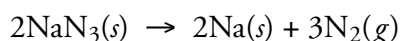
OBJECTIVE 19    **87.** Air bags in cars inflate when sodium azide, $NaN_3$, decomposes to generate nitrogen gas. How many grams of $NaN_3$ must react to generate 112 L $N_2$ at 121 kPa and 305 K?

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

OBJECTIVE 19    88. When natural gas is heated to 370 °C, the sulfur it contains is converted to hydrogen sulfide. The $H_2S$ gas can be removed by a reaction with sodium hydroxide that forms solid sodium sulfide and water.

$$H_2S + 2NaOH \rightarrow Na_2S + 2H_2O$$
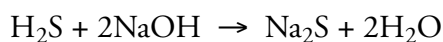
What volume of $H_2S(g)$ at 370 °C and 1.1 atm can be removed by 2.7 Mg NaOH?

OBJECTIVE 19    **89.** The hydrogen chloride gas used to make polyvinyl chloride (PVC) plastic is made by reacting hydrogen gas with chlorine gas. What volume of $HCl(g)$ at 105 kPa and 296 K can be formed from 1150 $m^3$ of $H_2(g)$ at STP?

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$
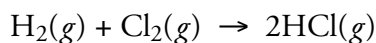
OBJECTIVE 19    90. The hydrogen gas needed to make ammonia, hydrogen chloride gas, and methanol can be obtained from small hydrocarbons such as methane or propane through the steam-reforming process, conducted at 41 atm and 760-980 °C. If $2.7 \times 10^7$ L of $C_3H_8(g)$ at 810 °C and 8.0 atm react in the first step of the process, shown below, what is the maximum volume in liters of $CO(g)$ at STP that can form?

$$C_3H_8 + 3H_2O \rightarrow 3CO + 7H_2$$
$$3CO + 3H_2O \rightarrow 3CO_2 + 3H_2$$
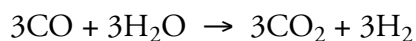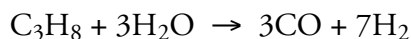
OBJECTIVE 19    **91.** Ammonia and carbon dioxide combine to produce urea, $NH_2CONH_2$, and water through the reaction shown below. About 90% of the urea is used to make fertilizers, but some is also used to make animal feed, plastics, and pharmaceuticals. To get the optimum yield of urea, the ratio of $NH_3$ to $CO_2$ is carefully regulated at a 3:1 molar ratio of $NH_3$ to $CO_2$.

$$2NH_3(g) + CO_2(g) \rightarrow NH_2CONH_2(s) + H_2O(l)$$

   a. What volume of $NH_3$ at STP must be combined with $1.4 \times 10^4$ L $CO_2$ at STP to yield the 3:1 molar ratio?

   b. The concentration of urea in a typical liquid fertilizer is about 1.0 M $NH_2CONH_2$. What minimum volume of $CO_2$ at 190 °C and 34.0 atm is needed to make $3.5 \times 10^4$ L of 1.0 M $NH_2CONH_2$?

OBJECTIVE 19    92. Alka-Seltzer contains the base sodium hydrogen carbonate, $NaHCO_3$, which reacts with hydrochloric acid in the stomach to yield sodium chloride, carbon dioxide, and water. What volume of $CO_2$ gas at 104.5 kPa and 311 K would be formed when 10.0 mL of 1.0 M HCl is mixed with an excess of $NaHCO_3$?

OBJECTIVE 19    **93.** Some chlorofluorocarbons, CFCs, are formed from carbon tetrachloride, $CCl_4$, which can be made from the methane in natural gas, as shown below. What minimum volume in cubic meters of methane gas at STP must be added to completely react $1.9 \times 10^6$ L of $Cl_2(g)$ at STP?

$$CH_4(g) + 4Cl_2(g) \rightarrow CCl_4(l) + 4HCl(g)$$
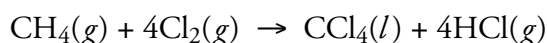
94. The carbon tetrachloride made in the the process described in the previous problem can be used to make CFC-12, $CF_2Cl_2$, and CFC-11, $CFCl_3$:

$$3HF(g) + 2CCl_4(g) \rightarrow CF_2Cl_2(g) + CFCl_3(g) + 3HCl(g)$$
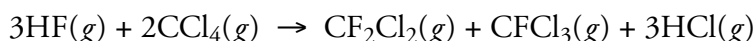
   a. What is the minimum volume of $HF(g)$ in cubic meters at 1.2 atm and 19 °C that would be necessary to react 235 kg $CCl_4$ in this reaction?
   b. What is the maximum volume of CFC-12 at 764 mmHg and 18 °C that can be formed from 724 m$^3$ of hydrogen fluoride gas at 1397 mmHg and 27 °C?

95. Ethylene oxide, $C_2H_4O$, is a cyclic compound whose two carbon atoms and oxygen atom form a 3-membered ring. The small ring is somewhat unstable, so ethylene oxide is a very reactive substance. It is useful as a rocket propellant, to sterilize medical plastic tubing, and to make polyesters and antifreeze. It is formed from the reaction of ethylene, $CH_2CH_2$, with oxygen at 270-290 °C and 8-20 atm in the presence of a silver catalyst.

$$2CH_2CH_2(g) + O_2(g) \rightarrow 2C_2H_4O(g)$$
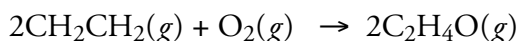
   a. What minimum mass in megagrams of liquefied ethylene, $CH_2CH_2$, is necessary to form $2.0 \times 10^5$ m$^3$ of ethylene oxide, $C_2H_4O$, at 107 kPa and 298 K?
   b. Ethylene can be purchased in cylinders containing 30 lb of liquefied ethylene. Assuming complete conversion of ethylene to ethylene oxide, how many of these cylinders would be necessary to make $2.0 \times 10^5$ m$^3$ of ethylene oxide at 107 kPa and 298 K?
   c. How many liters of ethylene oxide gas at 2.04 atm and 67 °C can be made from $1.4 \times 10^4$ L of $CH_2CH_2$ gas at 21 °C and 1.05 atm?
   d. How many kilograms of liquefied ethylene oxide can be made from the reaction of $1.4 \times 10^4$ L of $CH_2CH_2$ gas 1.05 atm and 21 °C and $8.9 \times 10^3$ L of oxygen gas at 0.998 atm and 19 °C?

96. Methanol, $CH_3OH$, a common solvent, is also used to make formaldehyde, acetic acid, and detergents. It is made from the methane in natural gas in two steps, the second of which is run at 200-300 °C and 50-100 atm using metal oxide catalysts:

$$3CH_4(g) + 2H_2O(g) + CO_2(g) \rightarrow 4CO(g) + 8H_2(g)$$
$$CO(g) + 2H_2(g) \rightarrow CH_3OH(l)$$
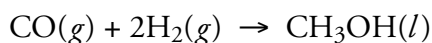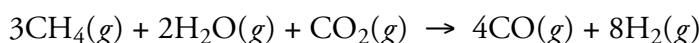
   a. In the second reaction, how many megagrams of liquid methanol can be produced from 187 m$^3$ of carbon monoxide, CO, at $2.0 \times 10^3$ kPa and 500 K?
   b. In the second reaction, what is the minimum volume of $H_2$ gas at $4.0 \times 10^3$ kPa and 500 K necessary to convert 187 m$^3$ of CO gas at $2.0 \times 10^3$ kPa and 500 K?
   c. How many megagrams of liquid methanol can be made from the reaction of 187 m$^3$ of $CO(g)$ at $2.0 \times 10^3$ kPa and 500 K and 159 m$^3$ of $H_2(g)$ at $4.0 \times 10^3$ kPa and 500 K?

OBJECTIVE 19

**97.** Rutile ore is about 95% titanium(IV) oxide, $TiO_2$, The $TiO_2$ is purified in the two reactions shown below. In the first reaction, what minimum volume in $m^3$ of $Cl_2$ at STP is necessary to convert all of the $TiO_2$ in $1.7 \times 10^4$ metric tons of rutile ore to $TiCl_4$?

$$3TiO_2 + 4C + 6Cl_2 \xrightarrow{900\ °C} 3TiCl_4 + 2CO + 2CO_2$$

$$TiCl_4 + O_2 \xrightarrow{1200\text{-}1400\ °C} TiO_2 + 2Cl_2$$
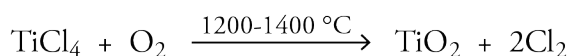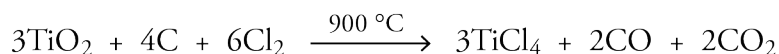
OBJECTIVE 19

98. Acetic acid, $CH_3CO_2H$, can be made from methanol and carbon dioxide in a reaction (shown below) that uses a rhodium/iodine catalyst. What is the percent yield when 513 kg of liquid $CH_3CO_2H$ form from the reaction of 309 kg $CH_3OH$ with 543 $m^3$ $CO(g)$ at 1 atm and 175 °C?

$$CH_3OH + CO \rightarrow CH_3CO_2H$$
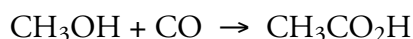
**Section 11.4 Dalton's Law of Partial Pressures**

OBJECTIVE 20

99. Explain why the total pressure of a mixture of ideal gases is equal to the sum of the partial pressures of each gas.

OBJECTIVE 21
OBJECTIVE 22

**100.** A typical 100-watt incandescent light bulb contains argon gas and nitrogen gas.
   a. If enough argon and nitrogen are added to this light bulb to yield a partial pressure of Ar of 98.5 kPa and a partial pressure of $N_2$ of 10.9 kPa, what is the total gas pressure in the bulb?
   b. If a 125-mL light bulb contains $5.99 \times 10^{-4}$ moles of $N_2$ and $5.39 \times 10^{-4}$ moles of Ar, what is the pressure in kPa in the bulb at a typical operating temperature of 119 °C?

OBJECTIVE 21
OBJECTIVE 22

101. A typical fluorescent light bulb contains argon gas and mercury vapor.
   a. If the total gas pressure in a fluorescent light bulb is 307.1 Pa and the partial pressure of the argon gas is 306.0 Pa, what is the partial pressure of the mercury vapor?
   b. If a 935-mL fluorescent bulb contains 4108 μg of argon gas and 77 μg of mercury vapor, what is the total gas pressure in the fluorescent bulb at 313 K?

**102.** If a 135-mL luminous tube contains 1.2 mg Ne and 1.2 mg Ar, at what temperature will the total gas pressure be 13 mmHg?

103. The atmosphere of Venus contains carbon dioxide and nitrogen gases. At the planet's surface, the temperature is about 730 K, the total atmospheric pressure is 98 atm, and the partial pressure of carbon dioxide is 94 atm. If scientists wanted to collect 10.0 moles of gas from the surface of Venus, what volume of gas should they collect?

**104.** Natural gas is a mixture of gaseous hydrocarbons and other gases in small quantities. The percentage of each component varies depending on the source. Consider a 21 $m^3$ storage container that holds 11.8 kg methane gas, $CH_4$, 2.3 kg ethane gas, $C_2H_6$, 1.1 kg propane gas, $C_3H_8$, and an unknown amount of other gases.
   a. If the total pressure in the container is 1.00 atm at 21 °C, how many moles of gases other than $CH_4$, $C_2H_6$, and $C_3H_8$ are in the container?
   b. What percentage of the gas particles in the cylinder are methane molecules?
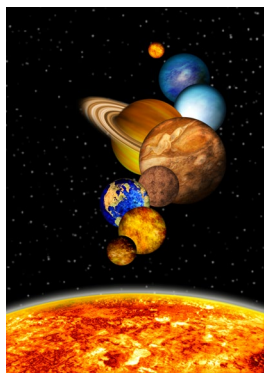
**Additional Problems**

105. One balloon is filled with helium, and a second balloon is filled to the same volume with xenon. Assuming that the two balloons are at the same temperature and pressure and that the gases are ideal, which of the following statements is true?
    a. The helium balloon contains more atoms than the xenon balloon.
    b. The helium balloon contains fewer atoms than the xenon balloon.
    c. The two balloons contain the same number of atoms.

**106.** Although the temperature decreases with increasing altitude in the troposphere (the lowest portion of the Earth's atmosphere), it increases again in the upper portion of the stratosphere. Thus, at about 50 km, the temperature is back to about 20 °C. Compare the following properties of the air along the California coastline at 20 °C and air at the same temperature but an altitude of 50 km. Assume that the percent composition of the air in both places is essentially the same.
    a. Do the particles in the air along the California coastline and the air at 50 km have the same average kinetic energy? If not, which has particles with the higher average kinetic energy? Explain your answer.
    b. Do the particles in the air along the California coastline and the air at 50 km have the same velocity? If not, which has particles with the higher average velocity? Explain your answer.
    c. Does the air along the California coastline and the air at 50 km have the same average distance between the particles? If not, which has the greater average distance between the particles? Explain your answer.
    d. Do the particles in the air along the California coastline and the air at 50 km have the same average frequency of particle collisions? If not, which has the greater frequency of between particle collisions? Explain your answer.
    e. Does the air along the California coastline and the air at 50 km have the same density? If not, which has the higher density? Explain your answer.
    f. Does the air along the California coastline and the air at 50 km have the same gas pressure? If not, which has the higher gas pressure? Explain your answer.

107. Compare the following properties of air on a sunny 28 °C day at a Florida beach to the properties of air along the coast of Alaska, where the temperature is –4 °C. Assume that the barometric pressure is the same in both places and that the percent composition of the air is essentially the same.
    a. Do the air in Florida at 28 °C and the air at –4 °C in Alaska have the same density? If not, which has the higher density?
    b. Do the air in Florida at 28 °C and the air at –4 °C in Alaska have the same average distance between the particles? If not, which has the higher average distance between the particles?

**108.** With reference to the relationships between properties of gases, answer the following questions.
    a. The pressure in the tires of a car sitting in a garage is higher at noon than at midnight. Why?
    b. The pressure in a car's tires increases as it is driven from home to school. Why?

109. With reference to the relationships between properties of gases, answer the following questions.
    a. When the air is pumped out of a gallon can, the can collapses. Why?
    b. A child's helium balloon escapes and rises into the air where it expands and pops. Why?

**110.** When people inhale helium gas and talk, their voices sound strange, as if it had been recorded and played back at a faster rate. Why do you think that happens?

111. If the following reaction is run in a cylinder with a movable piston that allows the pressure of gas in the cylinder to remain constant, and if the temperature is constant, what will happen to the volume of gas as the reaction proceeds? Why?

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

**Discussion Problems**

112. Consider two identical boxes, one containing helium gas and one containing xenon gas, both at the same temperature. Each box has a small hole in the side that allows atoms to escape. Which gas will escape faster? Why?

113. Is it easier to drink from a straw at the top of Mt Everest or at the seashore in Monterey, California? What causes the liquid to move up through a straw?

114. Like any other object that has mass, gas particles are attracted to the Earth by gravity. Why don't gas particles in a balloon settle to the bottom of the balloon?

115. When you look at a beam of sunlight coming through the window, you can see dancing particles of dust. Why are they moving? Why don't they settle to the floor?

116. Which of the following statements is/are true about a gas in a closed container?
    a. When the particles collide, the temperature rises in the container.
    b. When more gas is added to the container, the greater number of collisions between particles leads to an increase in temperature.
    c. If the gas is heated to a higher temperature, the number of collisions per second between particles will increase.

117. Baking powder contains sodium hydrogen carbonate, $NaHCO_3$, and a weak acid. Why does the reaction between these two components cause a donut to rise?

118. Use our model for gases to answer the following questions.
    a. How does a vacuum cleaner work?
    b. Why does a suction cup on a dart cause the dart to stick to a wall? Why does it help to wet the suction cup first?
    c. NASA is developing suction cup shoes for use in space stations. Why will they work inside the space station but not outside the space station?

119. To ensure that tennis balls in Denver will have the same bounce as tennis balls along the California coast, manufacturers put different gas pressures in them. Would you expect the gas pressure to be greater in a tennis ball intended for Denver or in one for San Francisco? Why?

120.  About 90% of the atoms in the universe are thought to be hydrogen atoms. Thus hydrogen accounts for about 75% of the mass of the universe. The Earth's crust, waters, and atmosphere contain about 0.9% hydrogen. Why is there so much less hydrogen on the Earth than in the rest of the universe? The planet Jupiter is thought to consist of about 92% hydrogen. Why is there so much more hydrogen on Jupiter than on Earth?

121.  When a scuba tank is filled from a storage tank, the scuba tank gets warmer and the temperature of the storage tank goes down. Why?

122.  Why does a helium filled balloon that escaped to the ceiling drop to the floor overnight?