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化学—中心科学

Chemistry—The Central Science

（英文版 · 原书第 8 版）

狄奥多尔　L. 布朗（Theodore　L. Brown）
（美）　小 H. 尤金 · 勒梅（H. Eugene LeMay, Jr.）　著
布鲁斯　E. 伯斯坦（Bruce　E. Bursten）

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出版说明

随着我国加入 WTO，国际间的竞争越来越激烈，而国际间的竞争实际上也就是人才的竞争、教育的竞争。为了加快培养具有国际竞争力的高水平技术人才，加快我国教育改革的步伐，国家教育部近来出台了一系列倡导高校开展双语教学、引进原版教材的政策。以此为契机，机械工业出版社拟于近期推出一系列国外影印版教材，其内容涉及高等学校公共基础课，以及机、电、信息领域的专业基础课和专业课。

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为了做好教材的引进工作，机械工业出版社特别成立了由著名专家组成的国外高校优秀教材审定委员会。这些专家对实施双语教学做了深入细致的调查研究，对引进原版教材提出许多建设性意见，并慎重地对每一本将要引进的原版教材一审再审，精选再精选，确认教材本身的质量水平，以及权威性和先进性，以期所引进的原版教材能适应我国学生的外语水平和学习特点。在引进工作中，审定委员会还结合我国高校教学课程体系的设置和要求，对原版教材的教学思想和方法的先进性、科学性严格把关。同时尽量考虑原版教材的系统性和经济性。

这套教材出版后，我们将根据各高校的双语教学计划，举办原版教材的教师培训，及时地将其推荐给各高校选用。希望高校师生在使用教材后及时反馈意见和建议，使我们更好地为教学改革服务。

机械工业出版社
2002 年 3 月

序

该书是一本已经再版7次的普通化学教材，三位作者都曾经多次获得过不同层次的教学奖。其中Brown在学术上造诣较深，曾获Guggenheim学者奖，美国化学会无机化学研究奖和无机化学进步服务优异奖等。现为Illinois大学（Urbana Champain）资深化学教授。LeMay教授有近30年的教学经验，并因此多次获得国家级教学奖。Bursten教授是Ohio大学杰出化学教授，在过渡金属和锕系元素化合物研究方面有较高的水平。从作者水平和该书的再版次数来看，这本书应当是质量比较好的。

该书图文并茂，插图精美。全书共分为25章，其内容安排和理论深度与国内现有的普通化学或大学一年级化学相近。应当认为，这类教材是20世纪70年代以来对我国普通化学或大学一年级化学影响最大的一类国外教材，比较重视化学基础理论的完整性和系统性。该书的初等量子力学和化学热力学部分写得比较适合大学一年级学生的水平。作为化学学科的学习，该书有较好的实用性。

该书设计了一些新的栏目，如关于解题（Problem Solving）设有：化学中实用的策略、例题、章后练习题（包括所附光盘上的e—媒体练习题）、中心科学演示（学生用CD—ROM和网站http：//www. prenhall. com/brown）；形象化（Visualization）设有：分子结构模型、化合物演示、分子运动三维动画以及中心科学演示；应用（Application）设有：生活和实际应用中的化学、化学新闻、有趣的阅读材料和中心科学演示。从而使教学内容可以方便地扩展到与生活和社会相关的其它方面。

该书可以用作双语基础化学教材。

宋心琦
清华大学化学系
2002年10月

Gases

10

10.1 Characteristics of Gases
10.2 Pressure
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10.5 Further Applications of the Ideal-Gas Equation
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10.7 Kinetic-Molecular Theory
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A tornado is one of the most destructive manifestations of the gaseous state of matter. Tornadoes form when moist, warm air at lower levels converges with cooler dry air aloft. The resultant air flows lead to winds that can attain speeds of 200 to 300 miles per hour.

In the past several chapters we have learned about electronic structures of atoms and about how atoms combine to form molecules and ionic substances. In everyday life, however, we encounter matter not on the atomic or molecular scale but as the large collections of atoms or molecules that we recognize as solids, liquids, and gases. ∞ (Section 1.2) This chapter will focus on gases; in Chapter 11 we will discuss liquids and solids.

We are surrounded by an atmosphere composed of a mixture of gases that we refer to as air. We breathe air to absorb oxygen, O_2, which supports human life. We also encounter gases in countless other situations. For example, chlorine gas, Cl_2, is used to purify drinking water. Acetylene gas, C_2H_2, is used in welding. Carbon dioxide, CO_2, and methane, CH_4, are among the gases known as *greenhouse gases*, which are implicated in global warming (Chapter 18). Although different gases may vary widely in their chemical properties, they share many physical properties. Our goals in this chapter are to gain a deeper understanding of the physical properties of gases and to learn what these properties teach us about the behavior of molecules.

10.1 Characteristics of Gases

Air consists primarily of oxygen and nitrogen. About 78 percent of the molecules in the air are N_2 molecules, and about 21 percent are O_2 molecules. Several other nonmetallic elements also exist as gases under ordinary conditions of temperature and pressure: H_2, F_2, Cl_2, and the noble gases He, Ne, Ar, Kr, and Xe. Many molecular compounds are also gases. Table 10.1 ▶lists a few of the more common gaseous compounds. Notice that all of these gases are composed entirely of nonmetallic elements. Furthermore, all have simple molecular formulas and, therefore, low molar masses.

Substances that are liquids or solids under ordinary conditions can usually also exist in the gaseous state, where they are often referred to as **vapors**. The substance H_2O, for example, can exist as liquid water, solid ice,

TABLE 10.1 Some Common Compounds That are Gases

Formula	Name	Characteristics
HCN	Hydrogen cyanide	Very toxic, slight odor of bitter almonds
HCl	Hydrogen chloride	Toxic, corrosive, choking odor
H_2S	Hydrogen sulfide	Very toxic, odor of rotten eggs
CO	Carbon monoxide	Toxic, colorless, odorless
CO_2	Carbon dioxide	Colorless, odorless
CH_4	Methane	Colorless, odorless, flammable
N_2O	Nitrous oxide	Colorless, sweet odor, laughing gas
NO_2	Nitrogen dioxide	Toxic, red-brown, irritating odor
NH_3	Ammonia	Colorless, pungent odor
SO_2	Sulfur dioxide	Colorless, irritating odor

or water vapor. Under the right conditions, a substance can coexist in all three states of matter, or *phases*, at the same time. A thermos bottle containing a mixture of ice and water at 0°C has some water vapor in the gas phase over the liquid and solid phases.

Gases differ significantly from solids and liquids in several respects. For example, a gas expands spontaneously to fill its container. Consequently, the volume of a gas equals the volume of the container in which it is held. Gases also are highly compressible: When pressure is applied to a gas, its volume readily decreases. In contrast, solids and liquids do not expand to fill their containers, and solids and liquids are not readily compressible.

Gases form homogeneous mixtures with each other regardless of the identities or relative proportions of the component gases. For example, when water and gasoline are poured into a bottle, the water vapor and gasoline vapors above the liquids form a homogeneous gas mixture. The two liquids, in contrast, do not mix; they remain as separate layers.

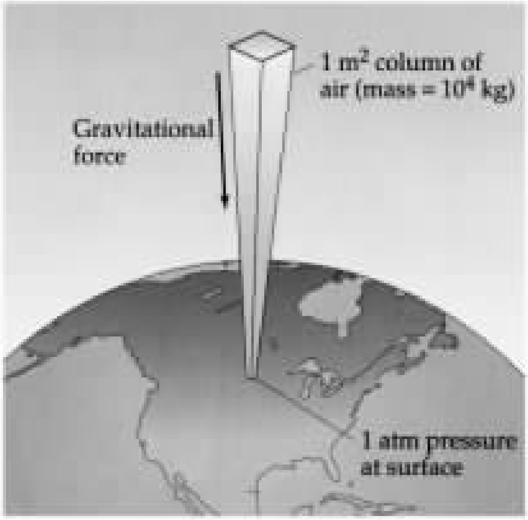
The characteristic properties of gases arise because the individual molecules are relatively far apart. For example, in the air we breathe, the molecules take up only about 0.1 percent of the total volume, with the rest being empty space. Thus, each molecule behaves largely as though the others weren't present. As a result, different gases behave similarly, even though they are made up of different molecules. In contrast, the individual molecules in a liquid are close together and occupy perhaps 70 percent of the total space. The attractive forces among the molecules keep the liquid together.

10.2 Pressure

Among the most readily measured properties of a gas are its temperature, volume, and pressure. It is not surprising, therefore, that many early studies of gases focused on relationships among these properties. We have already discussed volume and temperature. ∞ (Section 1.4) Let's now consider the concept of pressure.

In general terms, **pressure** conveys the idea of a force, a push that tends to move something else in a given direction. Pressure, *P*, is, in fact, the force, *F*, that acts on a given area, *A*:

$$P = \frac{F}{A} \qquad [10.1]$$



◀ **Figure 10.1** Illustration of the manner in which Earth's atmosphere exerts pressure at the surface of the planet. The mass of a column of atmosphere exactly 1 m^2 in cross-sectional area and extending to the top of the atmosphere exerts a force of 1.01×10^5 N.

Gases exert a pressure on any surface with which they are in contact. For example, the gas in an inflated balloon exerts a pressure on the inside surface of the balloon.

Atmospheric Pressure and the Barometer

Because of gravity, our atmosphere exerts a downward force and consequently a pressure on Earth's surface. The force F exerted by any object is the product of its mass, m, times its acceleration, a: $F = ma$. The acceleration produced by Earth's gravity is 9.8 m/s^2. A column of air 1 m^2 in cross section extending through the atmosphere has a mass of roughly 10,000 kg (Figure 10.1 ▲). The force exerted by this column is

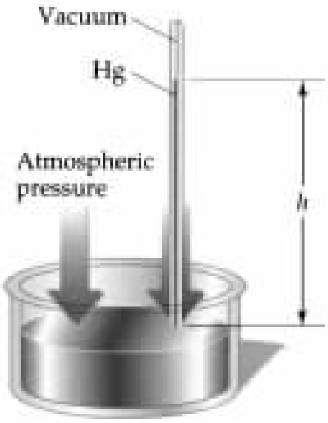
$$F = (10{,}000\ \text{kg})(9.8\ \text{m/s}^2) = 1 \times 10^5\ \text{kg-m/s}^2 = 1 \times 10^5\ \text{N}$$

The SI unit for force is kg-m/s^2 and is called the *newton* (N): 1 N = 1 kg-m/s^2. The pressure exerted by the column is the force divided by its cross-sectional area, A:

$$P = \frac{F}{A} = \frac{1 \times 10^5\ \text{N}}{1\ \text{m}^2} = 1 \times 10^5\ \text{N/m}^2 = 1 \times 10^5\ \text{Pa} = 1 \times 10^2\ \text{kPa}$$

The SI unit of pressure is N/m^2. It is given the name **pascal** (Pa) after Blaise Pascal (1623–1662), a French mathematician and scientist: 1 Pa = 1 N/m^2. A related unit sometimes used to report pressures is the **bar**, which equals 10^5 Pa. Atmospheric pressure at sea level is about 100 kPa or 1 bar. Of course, the actual atmospheric pressure at any location depends on weather conditions as well as on altitude.

▼ **Figure 10.2** Mercury barometer. The pressure of the atmosphere on the surface of the mercury (represented by the blue arrow) equals the pressure of the column of mercury (red arrow).



Atmospheric pressure can be measured by use of a mercury *barometer* (Figure 10.2 ▶). Such a barometer can be made from a glass tube more than 760 mm long that is closed at one end. The tube is completely filled with mercury and inverted into a dish that contains additional mercury. Care must be taken so that no air gets into the tube. Some of the mercury flows out when the tube is inverted, but a column of mercury remains in the tube. The space in the tube above the mercury is nearly a vacuum; only a negligible amount of mercury vapor occupies it.

The mercury surface in the dish experiences the full force, or weight, of Earth's atmosphere. Because there is no air (and therefore no atmospheric pressure) above the mercury in the tube, the mercury is pushed up the tube until the pressure due to the mass of the mercury column balances the atmospheric pressure. Thus, the height of the mercury column changes as the atmospheric pressure changes.

Standard atmospheric pressure, which corresponds to the typical pressure at sea level, is the pressure sufficient to support a column of mercury 760 mm in height. In SI units, this pressure equals 1.01325×10^5 Pa. It is used to define some common non-SI units used to express gas pressures: the **atmosphere** (atm) and the *millimeter of mercury* (mm Hg). The latter unit is also called the **torr**, after the Italian scientist Evangelista Torricelli (1608–1647), who invented the barometer:

$$1 \text{ atm} = 760 \text{ mm Hg} = 760 \text{ torr} = 1.01325 \times 10^5 \text{ Pa} = 101.325 \text{ kPa}$$

Note that the units mm Hg and torr are equivalent: 1 torr = 1 mm Hg.

We will usually express gas pressure in units of atm or torr, so you should be comfortable converting gas pressures from one set of units to another.

SAMPLE EXERCISE 10.1

(a) Convert 0.357 atm to torr. **(b)** Convert 6.6×10^{-2} torr to atm. **(c)** Convert 147.2 kPa to torr.

Solution In each case we are given the pressure in one unit and asked to convert it to another unit. In solving problems of this type, we can use dimensional analysis. The relationships given above for the various units of pressure are used to generate the conversion factors we need. **(a)** We convert atmospheres to torr by using the conversion factor derived from 760 torr = 1 atm:

$$(0.357 \cancel{\text{atm}})\left(\frac{760 \text{ torr}}{1 \cancel{\text{atm}}}\right) = 271 \text{ torr}$$

Note that the units cancel in the required manner.

(b) We use the same relationship that we did in part (a). Dimensional analysis tells us that we must use the conversion factor in the following way:

$$(6.6 \times 10^{-2} \cancel{\text{torr}})\left(\frac{1 \text{ atm}}{760 \cancel{\text{torr}}}\right) = 8.7 \times 10^{-5} \text{ atm}$$

(c) The relationship 760 torr = 101.325 kPa allows us to write an appropriate conversion factor for this problem:

$$(147.2 \text{ kPa})\left(\frac{760 \text{ torr}}{101.325 \text{ kPa}}\right) = 1104 \text{ torr}$$

PRACTICE EXERCISE

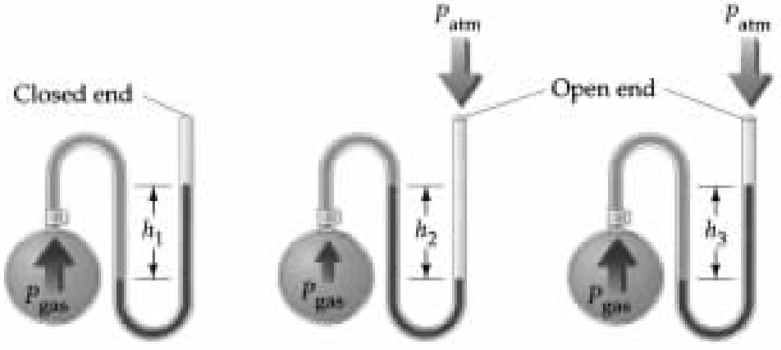
In countries that use the metric system—for example, Canada—atmospheric pressure in weather reports is given in units of kPa. Convert a pressure of 745 torr to kPa.
Answer: 99.3 kPa

▼ **Figure 10.3** A tire pressure gauge is used to measure the extent to which the pressure of the enclosed air exceeds atmospheric pressure.

Pressures of Enclosed Gases and Manometers

We use various devices to measure the pressures of enclosed gases. The tire gauge used to measure the pressure of air in automobile tires gives the pressure over and above atmospheric pressure (see Figure 10.3 ◀). In laboratories we sometimes use a device called a *manometer*, whose principle of operation is similar to that of a barometer. Figure 10.4(a) ▶ shows a closed-tube manometer, normally used to measure pressures below atmospheric pressure. The pressure equals the difference in the heights of the mercury levels in the two arms.

An open-tube manometer, like that pictured in Figure 10.4(b) and (c), is often employed to measure gas pressures that are near atmospheric pressure. The dif-



(a) $P_{gas} = P_{h_1}$ (b) $P_{gas} + P_{h_2} = P_{atm}$ (c) $P_{gas} = P_{atm} + P_{h_3}$

◀ **Figure 10.4** Closed-end manometer (a) and open-end manometers (b and c). In (b) gas pressure is less than atmospheric pressure; in (c) gas pressure exceeds atmospheric pressure.

ference in the heights of the mercury levels in the two arms of the manometer relates the gas pressure to atmospheric pressure. If the pressure of the enclosed gas is the same as atmospheric pressure, the levels in the two arms are equal. If the pressure of the enclosed gas is less than atmospheric pressure, the atmosphere will force the mercury higher in the arm exposed to the enclosed gas, as in Figure 10.4(b). In contrast, if the gas pressure exceeds atmospheric pressure, the mercury will be forced higher in the arm exposed to the atmosphere, as in Figure 10.4(c).

Although mercury is most often used in a manometer, other liquids can be used. For a given pressure difference, the difference in heights of the liquid levels in the two arms of the manometer is inversely proportional to the density of the liquid. The greater the density of the liquid, the smaller is the difference in column heights. The high density of mercury (13.6 g/mL) allows us to build smaller manometers than we could with less dense liquids.

SAMPLE EXERCISE 10.2

On a certain day the barometer in a laboratory indicates that the atmospheric pressure is 764.7 torr. A sample of gas is placed in a vessel attached to an open-end mercury manometer. A meter stick is used to measure the height of the mercury above the bottom of the manometer. The level of mercury in the open-end arm of the manometer has a measured height of 136.4 mm, and that in the arm that is in contact with the gas has a height of 103.8 mm. What is the pressure of the gas in atmospheres?

Solution We are given the atmospheric pressure (764.7 torr), and the fact that the mercury level in the arm of the manometer that is open to the atmosphere is higher (136.4 mm) than that in the arm attached to the enclosed gas (103.8 mm). Because the mercury in the open arm is higher than that in the enclosed arm, the pressure of the gas is greater than atmospheric pressure, as in Figure 10.4(c).

The difference in height between the two arms gives the amount by which the pressure of the gas exceeds atmospheric pressure. Because a mercury manometer is used, the height difference directly measures the pressure difference in mm Hg or torr. Thus, the pressure of the gas is equal to the atmospheric pressure plus the difference in height between the two arms of the manometer:

$$\begin{aligned} P_{gas} &= P_{atm} + (\text{difference in height of arms}) \\ &= 764.7\ \text{torr} + (136.4\ \text{torr} - 103.8\ \text{torr}) \\ &= 797.3\ \text{torr} \end{aligned}$$

Finally, we convert the pressure of the gas to atmospheres:

$$P_{gas} = (797.3\ \text{torr})\left(\frac{1\ \text{atm}}{760\ \text{torr}}\right) = 1.049\ \text{atm}$$

Note that in order to use an open-end manometer, we must know the value of the atmospheric pressure.

Chemistry and Life Blood Pressure

The human heart pumps blood to the parts of the body through arteries, and the blood returns to the heart through veins. When your blood pressure is measured, two values are reported, such as 120/80 (120 over 80), which is a normal reading. The first measurement is the *systolic pressure*, the maximum pressure when the heart is pumping. The second is the *diastolic pressure*, the pressure when the heart is in the resting part of its pumping cycle. The units associated with these pressure measurements are torr.

Blood pressure is measured using a pressure gauge attached to a closed, air-filled jacket or cuff that is applied like a tourniquet to the arm (Figure 10.5 ▶). The pressure gauge may be a mercury manometer or some other device. The air pressure in the cuff is increased using a small pump until it is above the systolic pressure and prevents the flow of blood. The air pressure inside the cuff is then slowly reduced until blood just begins to pulse through the artery, as detected by the use of a stethoscope. At this point the pressure in the cuff equals the pressure that the blood exerts inside the arteries. Reading the gauge gives the systolic pressure. The pressure in the cuff is then reduced further until the blood flows freely. The pressure at this point is the diastolic pressure.

◀ **Figure 10.5** Measuring blood pressure.

Hypertension is the presence of abnormally high blood pressure. The usual criterion for hypertension is a blood pressure greater than 140/90. Hypertension significantly increases the workload on the heart and also places a stress on the walls of the blood vessels throughout the body. These effects increase the risk of aneurysms, heart attacks, and strokes.

PRACTICE EXERCISE

A vessel connected to an open-end mercury manometer is filled with gas to a pressure of 0.835 atm. The atmospheric pressure is 755 torr. **(a)** In which arm of the manometer will the level of mercury be higher? **(b)** What is the height difference between the two arms of the manometer? *Answers:* **(a)** The level in the arm attached to the gas is higher. **(b)** 120 mm

10.3 The Gas Laws

Experiments with a large number of gases reveal that four variables are needed to define the physical condition, or *state*, of a gas: temperature, T, pressure, P, volume, V, and the amount of gas, which is usually expressed as the number of moles, n. The equations that express the relationships among T, P, V, and n are known as the *gas laws*.

The Pressure-Volume Relationship: Boyle's Law

If the pressure on a balloon is decreased, the balloon expands. That is why weather balloons expand as they rise through the atmosphere (Figure 10.6 ▶). Conversely, when a volume of gas is compressed, the pressure of the gas increases. British chemist Robert Boyle (1627–1691) first investigated the relationship between the pressure of a gas and its volume.

To perform his gas experiments, Boyle used a J-shaped tube like that shown in Figure 10.7 ▶. A quantity of gas is trapped in the tube behind a column of mercury. Boyle changed the pressure on the gas by adding mercury to the tube. He found that the volume of the gas decreased as the pressure increased. For

◀ **Figure 10.6** The volume of gas in this weather balloon will increase as it ascends into the high atmosphere, where the atmospheric pressure is lower than on Earth's surface.

example, doubling the pressure caused the gas volume to decrease to half its original value.

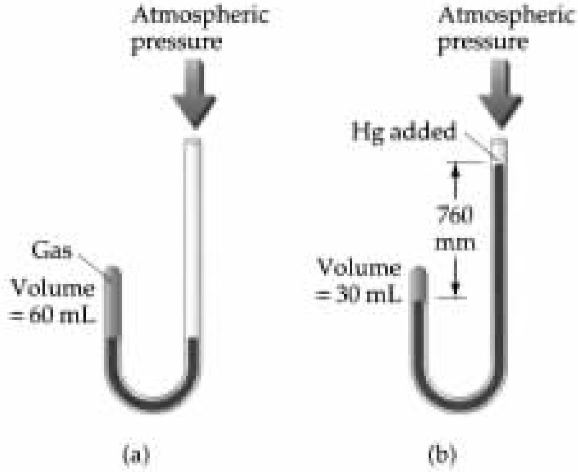
Boyle's law, which summarizes these observations, states that *the volume of a fixed quantity of gas maintained at constant temperature is inversely proportional to the pressure*. When two measurements are inversely proportional, one gets smaller as the other gets larger. Boyle's law can be expressed in mathematical terms:

P-V Relationships movie

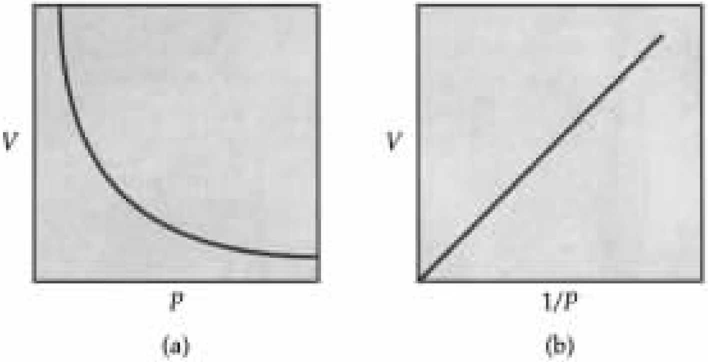
$$V = \text{constant} \times \frac{1}{P} \quad \text{or} \quad PV = \text{constant} \qquad [10.2]$$

The value of the constant depends on the temperature and the amount of gas in the sample. The graph of V versus P in Figure 10.8(a) ▶ shows the type of curve obtained for a given quantity of gas at a fixed temperature. A linear relationship is obtained when V is plotted versus $1/P$ [Figure 10.8(b)].

We see an application of Boyle's law every time we breathe. Between breaths the gas pressure inside the lungs equals atmospheric pressure. The volume of the lungs is governed by the rib cage, which can expand and contract, and the diaphragm, a muscle beneath the lungs. Inhalation occurs when the rib cage ex-



◀ **Figure 10.7** An illustration of Boyle's experiment. In (a) the volume of the gas trapped in the J-tube is 60 mL when the gas pressure is 760 torr. When additional mercury is added, as shown in (b), the trapped gas is compressed. The volume is 30 mL when its total pressure is 1520 torr, corresponding to atmospheric pressure plus the pressure exerted by the 760-mm column of mercury.



▶ **Figure 10.8** Graphs based on Boyle's law: (a) volume versus pressure; (b) volume versus 1/*P*.

pands and the diaphragm moves downward. Both of these actions serve to increase the volume of the lungs, thus decreasing the gas pressure inside the lungs. The atmospheric pressure then forces air into the lungs until the pressure in the lungs once again equals atmospheric pressure. Exhalation involves the reverse process: The rib cage contracts and the diaphragm moves up, both of which decrease the volume of the lungs. Air is forced out of the lungs by the increase in pressure caused by this reduction in volume.

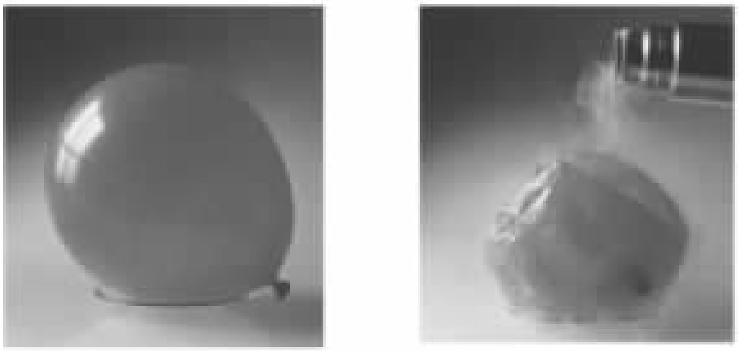
The Temperature-Volume Relationship: Charles's Law

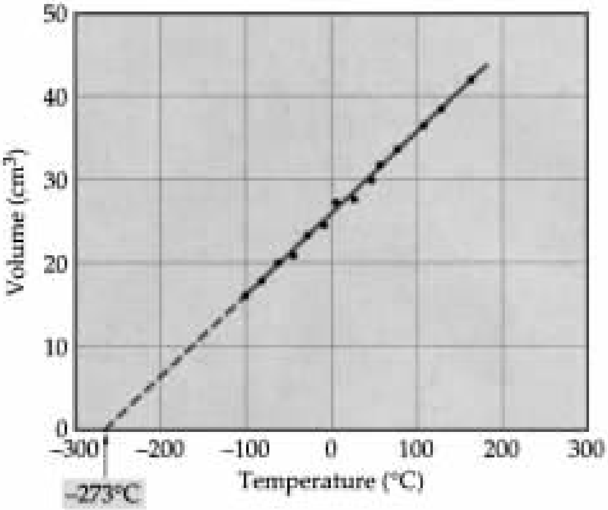
Hot-air balloons rise because air expands as it is heated. The warm air in the balloon is less dense than the surrounding cool air at the same pressure. This difference in density causes the balloon to ascend. Conversely, a balloon will shrink when the gas in it is cooled, as seen in Figure 10.9 ▼.

The relationship between gas volume and temperature was discovered in 1787 by the French scientist Jacques Charles (1746–1823). Charles found that the volume of a fixed quantity of gas at constant pressure increases linearly with temperature. Some typical data are shown in Figure 10.10 ▶. Notice that the extrapolated (extended) line (which is dashed) passes through −273°C. Note also that the gas is predicted to have zero volume at this temperature. Of course, this condition is never fulfilled because all gases liquefy or solidify before reaching this temperature.

In 1848 William Thomson (1824–1907), a British physicist whose title was Lord Kelvin, proposed an absolute-temperature scale, now known as the Kelvin scale. On this scale 0 K, which is called *absolute zero*, equals −273.15°C. ∞ (Section 1.4) In terms of the Kelvin scale, **Charles's law** can be stated as follows: *The volume of a fixed amount of gas maintained at constant pressure is directly proportional to its absolute temperature.* Thus, doubling the absolute temperature, say

▶ **Figure 10.9** As liquid nitrogen (−196°C) is poured over a balloon, the gas in the balloon is cooled and its volume decreases.





◀ **Figure 10.10** Volume of an enclosed gas as a function of temperature at constant pressure. The dashed line is an extrapolation to temperatures at which the substance is no longer a gas.

from 200 K to 400 K, causes the gas volume to double. Mathematically, Charles's law takes the following form:

$$V = \text{constant} \times T \qquad \text{or} \qquad \frac{V}{T} = \text{constant} \qquad [10.3]$$

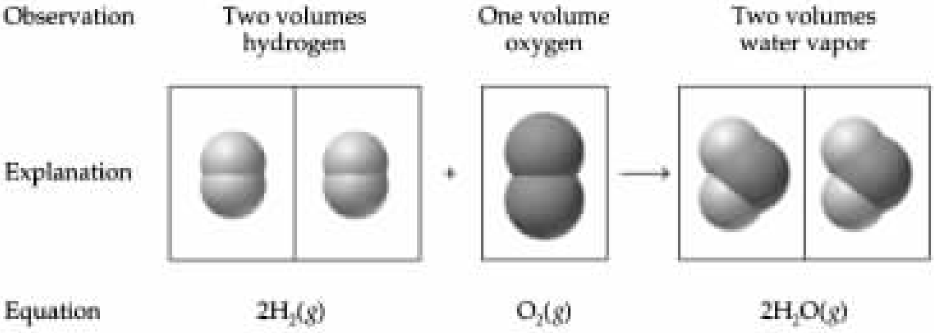
The value of the constant depends on the pressure and amount of gas.

The Quantity-Volume Relationship: Avogadro's Law

Gas Laws simulation

As we add gas to a balloon, the balloon expands. The volume of a gas is affected not only by pressure and temperature but by the amount of gas as well. The relationship between the quantity of a gas and its volume follows from the work of Joseph Louis Gay-Lussac (1778–1823) and Amadeo Avogadro.

Gay-Lussac is one of those extraordinary figures in the history of science who could truly be called an adventurer. He was interested in lighter-than-air balloons, and in 1804 he made an ascent to 23,000 ft—an exploit that set the altitude record for several decades. To better control lighter-than-air balloons, Gay-Lussac carried out several experiments on the properties of gases. In 1808 he observed the *law of combining volumes*: At a given pressure and temperature the volumes of gases that react with one another are in the ratios of small whole numbers. For example, two volumes of hydrogen gas react with one volume of oxygen gas to form two volumes of water vapor, as shown in Figure 10.11 ▼.



◀ **Figure 10.11** Gay-Lussac's experimental observation of combining volumes shown together with Avogadro's explanation of this phenomenon.

Volume	22.4 L	22.4 L	22.4 L
Pressure	1 atm	1 atm	1 atm
Temperature	0°C	0°C	0°C
Mass of gas	4.00 g	28.0 g	16.0 g
Number of gas molecules	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

▶ **Figure 10.12** Comparison illustrating Avogadro's hypothesis. Note that helium gas consists of helium atoms. Each gas has the same volume, temperature, and pressure and thus contains the same number of molecules. Because a molecule of one substance differs in mass from a molecule of another, the masses of gas in the three containers differ.

Three years later Amadeo Avogadro (Section 3.4) interpreted Gay-Lussac's observation by proposing what is now known as **Avogadro's hypothesis**: *Equal volumes of gases at the same temperature and pressure contain equal numbers of molecules*. For example, experiments show that approximately 22.4 L of any gas at 0°C and 1 atm contain 6.02×10^{23} gas molecules (that is, 1 mol), as depicted in Figure 10.12 ▲.

Avogadro's law follows from Avogadro's hypothesis: *The volume of a gas maintained at constant temperature and pressure is directly proportional to the number of moles of the gas*. That is,

$$V = \text{constant} \times n \qquad [10.4]$$

Thus, doubling the number of moles of gas will cause the volume to double if T and P remain constant.

10.4 The Ideal-Gas Equation

In the preceding section we examined three historically important gas laws. Each was obtained by holding two variables constant in order to see how the other variables affect each other. Using the symbol $\propto$, which is read "is proportional to," we have

$$\text{Boyle's law:}\quad V \propto \frac{1}{P} \qquad (\text{constant } n, T)$$

$$\text{Charles's law:}\quad V \propto T \qquad (\text{constant } n, P)$$

$$\text{Avogadro's law:}\quad V \propto n \qquad (\text{constant } P, T)$$

We can combine these relationships to make a more general gas law:

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

If we call the proportionality constant R, we obtain

$$V = R\left(\frac{nT}{P}\right)$$

Rearranging, we have this relationship in its more familiar form:

$$PV = nRT \qquad [10.5]$$

This equation is known as the **ideal-gas equation**. An **ideal gas** is a hypothetical gas whose pressure, volume, and temperature behavior is completely described by the ideal-gas equation.

The term R in the ideal-gas equation is called the **gas constant**. The value and units of R depend on the units of P, V, n, and T. Temperature must *always* be expressed as an absolute temperature. The quantity of gas, n, is normally expressed in moles. The units chosen for pressure and volume are most often atm and liters, respectively. However, other units can be used. Table 10.2 ▶ shows the numerical value for R in various units. As we saw in the "Closer Look" box in Section 5.3, the product PV has the units of energy. Therefore, the units of R can include calories or joules. In working problems with the ideal-gas equation, the units of P, V, n, and T must agree with the units in the gas constant. In this chapter we will use the value $R = 0.08206$ L-atm/mol-K (four significant figures) or 0.0821 L-atm/mol-K (three significant figures) whenever we use the ideal-gas equation.

TABLE 10.2 Numerical Values of the Gas Constant, *R*, in Various Units

Units	Numerical Value
L-atm/mol-K	0.08206
cal/mol-K	1.987
J/mol-K[a]	8.314
m^3-Pa/mol-K[a]	8.314
L-torr/mol-K	62.36

[a] SI unit.

Suppose we have 1.000 mol of an ideal gas at 1.000 atm and 0.00°C (273.15 K). Then, from the ideal-gas equation the volume of the gas is:

$$V = \frac{nRT}{P} = \frac{(1.000\ \text{mol})(0.08206\ \text{L-atm/mol-K})(273.15\ \text{K})}{1.000\ \text{atm}} = 22.41\ \text{L}$$

The conditions 0°C and 1 atm are referred to as the **standard temperature and pressure (STP)**. Many properties of gases are tabulated for these conditions. The volume occupied by 1 mol of ideal gas at STP, 22.41 L, is known as the *molar volume* of an ideal gas at STP.

The ideal-gas equation does not always accurately describe real gases. Thus, the measured volume, V, for given conditions of P, n, and T, might differ from the volume calculated from $PV = nRT$. For example, the measured molar volumes of real gases at STP differ slightly from that calculated for an ideal gas, as shown in Figure 10.13 ▶. Ordinarily the difference between ideal and real behavior is so small that we can ignore it.

SAMPLE EXERCISE 10.3

Calcium carbonate, $CaCO_3(s)$, decomposes upon heating to give $CaO(s)$ and $CO_2(g)$. A sample of $CaCO_3$ is decomposed, and the carbon dioxide is collected in a 250-mL flask. After the decomposition is complete, the gas has a pressure of 1.3 atm at a temperature of 31°C. How many moles of CO_2 gas were generated?

Solution We are given the volume (250 mL), pressure (1.3 atm), and temperature (31°C) of a sample of CO_2 gas and asked to calculate the number of moles of CO_2 in the sample. Because we are given V, P, and T, we can solve the ideal gas equation for the unknown quantity, n.

In analyzing and solving gas-law problems, it is helpful to tabulate the information given in the problems and then to convert the values to units that are consistent with those for R (0.0821 L-atm/mol-K). In this case the given values are

$$P = 1.3\ \text{atm}$$
$$V = 250\ \text{mL} = 0.250\ \text{L}$$
$$T = 31^\circ\text{C} = (31 + 273)\ \text{K} = 304\ \text{K}$$

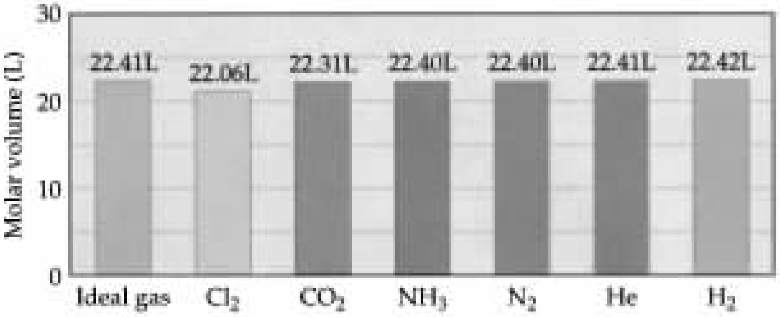
Remember: *Absolute temperature must always be used when the ideal-gas equation is solved.*

We now rearrange the ideal-gas equation (Equation 10.5) to solve for n.

$$n = \frac{PV}{RT}$$

$$n = \frac{(1.3\ \text{atm})(0.250\ \text{L})}{(0.0821\ \text{L-atm/mol-K})(304\ \text{K})} = 0.013\ \text{mol}\ CO_2$$

Notice that appropriate cancellation of units ensures that we have properly rearranged the ideal-gas equation and have converted to the correct units.



▶ **Figure 10.13** Comparing the molar volumes at STP of the ideal gas with various real gases.

PRACTICE EXERCISE

Tennis balls are usually filled with air or N_2 gas to a pressure above atmospheric pressure to increase their "bounce." If a particular tennis ball has a volume of 144 cm^3 and contains 0.33 g of N_2 gas, what is the pressure inside the ball at 24°C? ***Answer:*** 2.0 atm

Strategies in Chemistry Calculations Involving Many Variables

In chemistry and throughout your studies of science and math you may encounter problems that involve several experimentally measured variables as well as several different physical constants. In this chapter we encounter a variety of problems based on the ideal-gas equation. That equation involves four experimental quantities—P, V, n, and T—and one constant, R. Depending on the type of problem, we might need to solve for any of the four quantities.

To avoid any difficulty extracting the necessary information from problems when many variables are involved, we suggest you take the following steps as you analyze, plan, and solve such problems:

1. *Tabulate information.* Read the problems carefully to determine which quantity is the unknown and which quantities are given. Every time you encounter a numerical value, jot it down. In many cases, constructing a table of the given information will be useful.

2. *Convert to appropriate units.* As you have already seen, we often use several different units to express the same quantity. You must make certain that quantities are converted to the proper units by using the correct conversion factors. In using the ideal-gas equation, for example, we usually use the value of R that has units of L-atm/mol-K. If you are given a pressure in torr, you will need to convert it to atmospheres.

3. *If a single equation relates the variables, rearrange the equation to solve for the unknown.* You should make certain that you are comfortable using algebra to solve the equation for the desired variable. In the case of the ideal-gas equation, the following algebraic rearrangements will all be used at one time or another.

$$P = \frac{nRT}{V}; \quad V = \frac{nRT}{P}; \quad n = \frac{PV}{RT}; \quad T = \frac{PV}{nR}$$

4. *Use dimensional analysis:* Carry the units through your calculation. Use of dimensional analysis enables you to check that you have solved the equation correctly. If the units of the quantities in the equation cancel properly to give the units of the desired variable, you have probably used the equation correctly.

One note of caution is in order: Sometimes you will not be given values for the necessary variables directly. Rather, you will be given the values of other quantities that can be used to determine the needed variables. For example, suppose you are trying to use the ideal-gas equation to calculate the pressure of a gas. You are given the temperature of the gas, but you are not given explicit values for n and V. However, the problem states that "the sample of gas contains 0.15 mol of gas per liter." We turn this statement into the expression

$$\frac{n}{V} = 0.15 \text{ mol/L}$$

Solving the ideal-gas equation for pressure yields

$$P = \frac{nRT}{V} = \left(\frac{n}{V}\right)RT$$

Thus, we can solve the equation even though we are not given specific values for n and V. We will examine how to use the density and molar mass of a gas in this fashion in Section 10.5.

As we have continuously stressed, the most important thing you can do to become proficient at solving problems is to practice by using practice exercises and assigned exercises at the end of each chapter. By using systematic procedures, such as those described here, you should be able to minimize difficulties in solving problems involving many variables.