

Chapter 13 Gases

13.1 Describing the Properties of Gases

Key Terms

barometer	torr (mm Hg)	standard atmosphere
pascal	Boyle's law	absolute zero
Charles's law	Avogadro's law	

Summary

Gases exert pressure on their surroundings. A *barometer* measures atmospheric pressure. In a barometer, the weight of the atmosphere presses on mercury in a dish, forcing the mercury up a thin tube. At sea level, the height of the column of mercury averages 760 mm. This measurement results in a commonly used unit of measurement for atmospheric pressure, millimeters of mercury (*mm Hg*), or the *torr*. Another unit of measurement is the *standard atmosphere* (atm). The SI unit for pressure is the *pascal* (Pa). The relationship among these units is as follows:

$$1.000 \text{ atm} = 760.0 \text{ mm Hg} = 760.0 \text{ torr} = 101,325 \text{ Pa}$$

As the volume of a trapped gas increases, its pressure decreases (at constant temperature and moles of gas). Pressure and volume are inversely proportional. This relationship between pressure and volume can be expressed as an equation, $PV = k$, which means that pressure times volume equals a constant value. This relationship is known as *Boyle's law*.

Charles's law is expressed by the equation $V = bT$, which means that for given sample of gas at constant pressure, volume equals temperature in kelvins times a constant value. Volume and temperature (in kelvins) are directly proportional. Charles's law suggests that every type of gas reaches a volume of zero when it reaches a certain temperature, -273°C , which is defined as *absolute zero* on the Kelvin scale.

Avogadro's law states that the volume of a gas is directly proportional to the number of moles if temperature and pressure remain constant. This relationship can be expressed by the following equations: $V = an$ or $V/n = a$.

13.2 Using Gas Laws to Solve Problems

Key Terms

universal gas constant	ideal gas law	ideal gas
combined gas law	partial pressures	molar volume
Dalton's law of partial pressures	standard temperature and pressure (STP)	

Summary

Boyle's law, Charles's law, and Avogadro's law can be combined into the *ideal gas law*, or $PV = nRT$, where P stands for pressure, V for volume, n for number of moles, T for temperature, and R stands for the *universal gas constant*, which combines the constants of the three individual gas laws. It is equal to $0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol}$, or $8.314 \text{ J/K} \cdot \text{mol}$.

Dalton's law of partial pressures states that for a mixture of gases in a container, the total pressure exerted is the sum of the *partial pressures* (the pressures that individual gases would exert if they were alone in a container) of the gases present.

The ideal gas equation makes it possible to calculate the number of moles present in a sample if its pressure, volume, and temperature are known. The *molar volume* of an ideal gas is 22.4 L at 0°C (273 K) and 1 atm . The conditions 0°C and 1 atm are called *standard temperature and pressure* (STP).

13.3 Using a Model to Describe Gases

Key Terms

kinetic molecular theory

Summary

A law is a general statement about behavior that has been observed in many experiments. Laws do not explain why nature behaves the way it does. Theories, or models, attempt to answer this question. Models in chemistry are explanations about how individual atoms or molecules cause the behavior of substances that can be observed.

The *kinetic molecular theory* tries to explain the behavior of an ideal gas. This theory is based on certain postulates, or assumptions. The postulates best explain the behavior of real gases at high temperatures and low pressures.

Postulates Supporting the Kinetic Molecular Theory of Ideal Gas Behavior

1	2	3	4	5
Gases are made up of tiny particles.	These particles are so small compared with the distances between them that the volume of the individual particles can be treated as zero.	The particles are in constant random motion. Their collisions with the walls of the container cause the pressure of the gas.	The particles neither attract nor repel (drive away) one another.	As the temperature increases, the motion of the particles (and therefore their kinetic energy) increases.

Heating a gas to a higher temperature causes its particles to move faster and hit the walls of its container more often and more forcefully. As a result, pressure increases as temperature increases. If the container is flexible, the gas's volume increases as a result.

Under conditions of high pressure and small volume, real gases behave differently from ideal gases and do not obey the ideal gas law ($PV = nRT$) very well.

Additional Active Reading Questions

1. What is the value in atmospheres of 770. mm Hg?
2. State Boyle's law.
3. What is significant about the temperature -273°C ?
4. What is the formula that expresses the ideal gas law? Which three gas laws does it combine?
5. Compute the total pressure for the following mixture of gases: $\text{N}_2 = 0.12 \text{ atm}$, $\text{CO}_2 = 0.70 \text{ atm}$.
6. Describe the postulate of the kinetic molecular theory that relates the average kinetic energy of a gas to its temperature.
7. What change causes the volume and the pressure of gases to increase?
8. What is the volume of 1 mol of an ideal gas at STP?