Gases and Kinetic Theory

- Ideal Gases
- The Kinetic Theory of Gases

Among the different phases of matter, the gaseous phase is the simplest to understand and to model, since all gases display similar behavior and follow similar laws regardless of their identity. The atoms or molecules in a gaseous sample move rapidly and are far apart. In addition, intermolecular forces between gas particles tend to be weak; this results in certain characteristic physical properties, such as the ability to expand to fill any volume and to take on the shape of a container. Furthermore, gases are easily, though not infinitely, compressible.

The state of a gaseous sample is generally defined by four variables: pressure (P), volume (V), temperature (T), and number of moles (n), though as we shall see, these are not all independent. The *pressure* of a gas is the force per unit area that the atoms or molecules exert on the walls of the container through collisions. The SI unit for pressure is the *pascal* (Pa), which is equal to one newton per meter squared. Sometimes gas pressures are expressed in *atmospheres* (atm). One atmosphere is equal to 10^5 Pa, and is approximately equal to the pressure the Earth's atmosphere exerts on us each day. Volume can be expressed in liters (L) or cubic meters (m^3), and temperature is measured in Kelvins (K) for the purpose of the gas laws. Recall that we can find the temperature in K by adding 273 to the temperature in Celsius. Gases are often discussed in terms of *standard temperature and pressure* (STP), which refers to the conditions of 273 K (0°C) and 1 atm.

The state of a gas can be defined by pressure, volume, temperature, and amount of gas in moles.

Ideal Gases

When examining the behavior of gases under varying conditions of temperature and pressure, it is most convenient to treat them as ideal gases. An ideal gas represents a hypothetical gas whose molecules have no intermolecular forces; that is, they do not interact with each other and occupy no volume. Although gases in reality deviate from this idealized behavior, at relatively low pressures and high temperatures many gases behave in nearly ideal fashion. Therefore, the assumptions used for ideal gases can be applied to real gases with reasonable accuracy.

> Most gases can be treated as ideal.

THE IDEAL GAS LAW

The ideal gas law gives the relationship between the pressure, volume, and temperature of a gas before and after some process. It states that the product of pressure and volume divided by temperature remains constant during any process:

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

The *ideal* gas law states that pressure times volume divided by temperature remains constant.

If temperature remains constant during a process (*isothermic*), the equation becomes

$$P_1 V_1 = P_2 V_2 \text{ (Boyle's law).}$$

If the volume remains constant during a process (*isochoric*), the equation becomes

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}\,.$$

If the pressure remains constant during a process (*isobaric*), the equation becomes

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ (Charles's law)}.$$

Example: Under isothermal conditions, what would be the volume of a 1 liter sample of helium gas after its pressure is changed from 12 atm to 4 atm?

Solution: Isothermal means that the temperature is constant during the process, so the ideal gas law becomes

$$P_1V_1 = P_2V_2$$

$$(12 \text{ atm})(1 \text{ L}) = (4 \text{ atm})V_2$$

$$V_2 = 3 \text{ L}.$$

Example: If the temperature of 2 liters of gas is isobarically changed from 10°C to 293°C, what would be the final volume?

Solution: *Isobarically* means that the pressure remains constant during this process. But before we can apply the ideal gas law, or Charles's law in this case, we need to convert Celsius degrees to Kelvins:

$$T_1 = 10°\text{C} + 273 = 283 \text{ K and } T_2 = 293°\text{C} + 273 = 566 \text{ K}$$

Then,

$$\begin{aligned} \frac{V_1}{T_1} &= \frac{V_2}{T_2} \\ \frac{2\text{ L}}{283\text{ K}} &= \frac{V_2}{566\text{ K}} \\ V_2 &= 4\text{ L}. \end{aligned}$$

QUICK QUIZ

True or False: As the temperature of a gas is doubled, the average speed of each molecule is also doubled.

Answer:

False. Temperature is proportional to the average kinetic energy of the molecules, and kinetic energy is proportional to the square of the speed of the molecules. Therefore, the temperature is proportional to the *square* of the average speed, not the average speed.