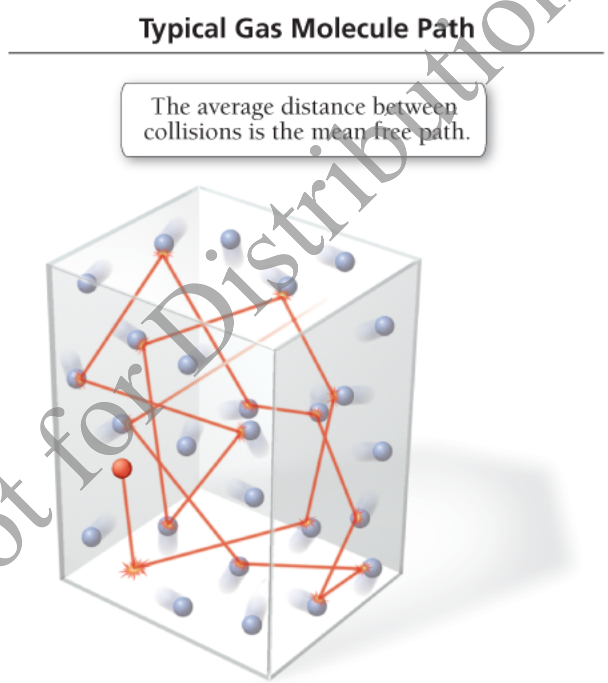
Pearson

10.9: Mean Free Path, Diffusion, and Effusion of Gases

We saw in Section 10.8 that the root mean square velocity of gas molecules at room temperature is in the range of hundreds of meters per second. However, suppose that your roommate just put on too much perfume in the bathroom only 2 m away. Why does it take a minute or two before you can smell the fragrance? Although most molecules in a perfume bottle have higher molar masses than nitrogen, their velocities are still hundreds of meters per second, so why the delay? The answer is that, even though gaseous particles travel at tremendous speeds, they also travel in haphazard paths (Figure 10.21). To a perfume molecule, the path from the perfume bottle in the bathroom to your nose 2 m away is like a bargain hunter's path through a busy shopping mall during a clearance sale. The molecule travels only a short distance before it collides with another molecule, changes direction, only to collide again, and so on. In fact, at room temperature and atmospheric pressure, a molecule in the air experiences several billion collisions per second.

Figure 10.21 Mean Free Path

A molecule in a volume of gas follows a haphazard path, involving many collisions with other molecules.



The average distance that a molecule travels between collisions is its **mean free path**. At room temperature and atmospheric pressure, the mean free path of a nitrogen molecule, which has a molecular diameter of 300 pm (four times the covalent radius), is 93 nm, or about 310 molecular diameters. If a nitrogen molecule were the size of a golf ball, it would travel about 40 ft between collisions. Mean free path increases with *decreasing* pressure. Under conditions of ultrahigh vacuum (10^{-10} torr), the mean free path of a nitrogen molecule is hundreds of kilometers.

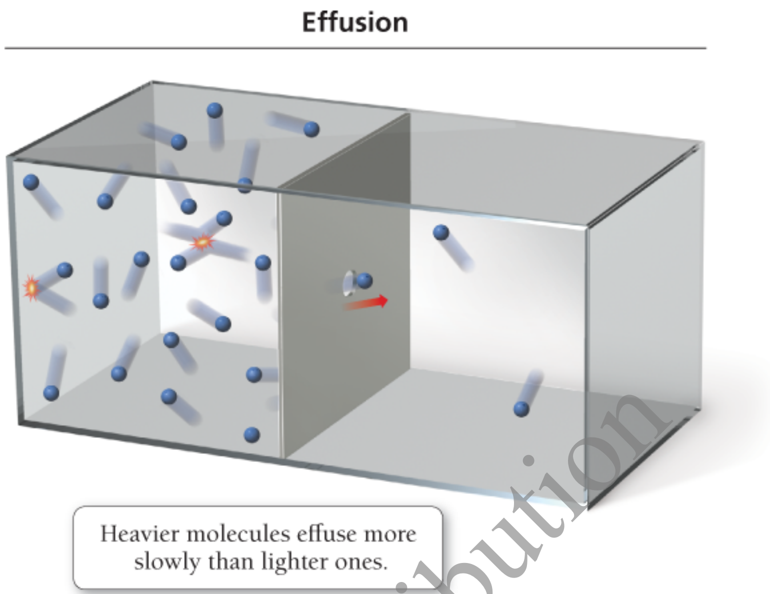
The process by which gas molecules spread out in response to a concentration gradient is **diffusion**, and even though the particles undergo many collisions, the root mean square velocity still influences the rate of diffusion. Heavier molecules diffuse more slowly than lighter ones, so the first molecules you smell from a perfume mixture (in a room with no air currents) are the lighter ones.

A process related to diffusion is **effusion**, the process by which a gas escapes from a container into a vacuum through a small hole (Figure 10.22). The rate of effusion is also related to root mean square velocity—heavier molecules effuse more slowly than lighter ones. The rate of effusion—the amount of gas that effuses in a given time—is inversely proportional to the square root of the molar mass of the gas, as follows:

$$\text{rate} \propto \frac{1}{\sqrt{M}}$$

Figure 10.22 Effusion

Effusion is the escape of a gas from a container into a vacuum through a small hole.



In a ventilated room, air currents greatly enhance the transport of gas molecules.

The ratio of effusion rates of two different gases is given by **Graham's law of effusion**, named after Thomas Graham (1805–1869):

[10.27]

$$\frac{\text{rate}_A}{\text{rate}_B} = \sqrt{\frac{M_B}{M_A}}$$

In this expression, rate_A and rate_B are the effusion rates of gases A and B and M_A and M_B are their molar masses.

Graham's law explains, in part, why helium balloons only float for a day or so. Because helium has such a low molar mass, it escapes from the balloon quite quickly. A balloon filled with air, by contrast, remains inflated longer because the gas particles within it have a higher average molar mass.

Example 10.13 Graham's Law of Effusion

An unknown gas effuses at a rate that is 0.462 times that of nitrogen gas (at the same temperature). Calculate the molar mass of the unknown gas in g/mol.

SORT The problem gives you the ratio of effusion rates for the unknown gas and nitrogen and asks you to find the molar mass of the unknown gas.

GIVEN:

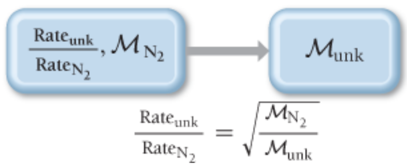
Rate

$$\frac{\text{Rate}_{\text{unk}}}{\text{Rate}_{\text{N}_2}} = 0.462$$

FIND: M_{nuk}

STRATEGIZE The conceptual plan uses Graham's law of effusion. You are given the ratio of rates, and you know the molar mass of the nitrogen. You can use Graham's law to determine the molar mass of the unknown gas.

CONCEPTUAL PLAN



RELATIONSHIP USED

$$\frac{\text{rate}_{\text{A}}}{\text{rate}_{\text{B}}} = \sqrt{\frac{M_{\text{B}}}{M_{\text{A}}}} \text{ (Graham's law)}$$

SOLVE Solve the equation for M_{nuk}, substitute the correct values, and calculate.

SOLUTION

$$\begin{aligned} \frac{\text{rate}_{\text{unk}}}{\text{rate}_{\text{N}_2}} &= \sqrt{\frac{M_{\text{N}_2}}{M_{\text{unk}}}} \\ M_{\text{nuk}} &= \frac{M_{\text{N}_2}}{\left(\frac{\text{rate}_{\text{unk}}}{\text{rate}_{\text{N}_2}}\right)^2} \\ &= \frac{28.02\text{ g/mol}}{(0.462)^2} = 131\text{ g/mol} \end{aligned}$$

CHECK The units of the answer are correct. The magnitude of the answer seems reasonable for the molar mass of a gas. In fact, from the answer, you can even conclude that the gas is probably xenon, which has a molar mass of 131.29 g/mol.

FOR PRACTICE 10.13 Find the ratio of effusion rates of hydrogen gas and krypton gas.

Interactive Worked Example 10.13 **Graham's Law of Effusion**