Appendix B
Gases

B.1 GAS MOLECULES ACTING COLLECTIVELY

According to the Kinetic Molecular Theory (CAMS Section 7.2), gases are in constant random motion and the average kinetic energy is proportional to the absolute temperature. The kinetic energy of a molecule is 1/2 mv^2, so the average speed of the molecules also depends on the absolute temperature. The average O_2 molecule moves at about 1,000 mph on a nice day.

However, it is the collective action of large numbers of molecules that we sense or measure as gases not the individual molecules. When you fan your face, you feel some wind, which is the effect of molecules in the air hitting your face. You cannot sense the individual molecules hitting your skin, for they are much too small, but you can feel their collective action.

The fact that a balloon expands when it is filled with a gas also shows how gas molecules act collectively. The molecules in the balloon are moving around with an average kinetic energy dictated by the temperature. When a molecule strikes the inside wall of the balloon, it exerts a force on the balloon and pushes it outward. The collective forces of all of the molecules inside the balloon pushing outward cause it to stay inflated. At the same time, the gas molecules in the outside air are striking the outer surface of the balloon exerting a force pushing inward. The size of the balloon adjusts until the force from the “strikes” on the outside balances the force from the “strikes” on the inside.

The collective force of all of the molecules pushing on the inside wall of the balloon results in pressure. The collective force per unit area or pressure of the gas depends on the number of collisions with the walls per second and the force of each collision. The common units are pounds per square inch, atmosphere (atm), the millimeter of mercury (mm Hg or torr) and the pascal, the SI unit of pressure (N/m^2).

B.2 RELATIONSHIP OF PRESSURE TO OTHER GAS PROPERTIES

Let’s analyze what happens to the pressure of a gas as the temperature, the number of molecules and volume of the gas change. Imagine a cylinder with a movable piston. The gas molecules in the piston have kinetic energy (are moving) and are hitting the walls of the cylinder, the piston and each other.

If the temperature of the gas is increased, the molecules will move faster and will strike the piston more frequently and with more force. Consequently the pressure increases. If more molecules are added to the cylinder (moles of gas increase), the frequency of collisions and therefore the pressure increases. Finally, if we push the piston down and compress the gas to a smaller volume, the gas molecules have less distance to travel before they hit the piston, and they collide with the piston more frequently. Thus, a decrease in volume will result in an increase in pressure.

The relationships among the pressure, volume, number of moles, and temperature of a gas are summed up quantitatively in the *ideal gas law*:

$$PV = nRT$$

P is the pressure in atmospheres (atm), V the volume in liters (L), n the number of moles, T the absolute temperature in kelving, and R is a constant called the ideal gas constant, which is 0.0821 $L{\cdot}atm{\cdot}K^{-1}{\cdot}mol^{-1}$. However, when using SI units, P is expressed in pascals, V in m^3, and R = 8.314 $J{\cdot}mol^{-1}{\cdot}K^{-1}$.

B.3 USING THE IDEAL GAS LAW

The ideal gas law contains four experimental quantities: pressure, volume, temperature, and number of moles. If we know three of the quantities, we can solve for the fourth. The first step is always to ensure that the units on the known quantities are consistent with our value of R. The following examples show how this can be done.

Example 1

What is the volume of 1.00 mole of gas at 1.00 atm and 0 ºC?

Solution:

n is in mol and P in atm, but T is in ºC, not K. To convert from ºC to K, we add 273: 0 ºC + 273 = 273 K = T. Next, rearrange the ideal gas law to solve for the unknown, which in this case is the volume.

$$\mathbf{V = \frac{nRT}{P} = \frac{(1.00\ mol)\ (0.0821\ L \cdot atm \cdot K^{-1} \cdot mol^{-1})\ (273\ K)}{1.00\ atm} = 22.4\ L}$$

Comment:

The conditions 0 ºC and 1.00 atm are often referred to as the Standard Temperature and Pressure (STP) for a gas. The volume at STP is 22.4 L, which is an experimental, not theoretical, number that students often remember from high school chemistry.

Example 2

An experiment yields 5.67 mL of CO_2(g) at 26 ºC and 782 mm Hg. How many grams of CO_2 is this?

Solution:

We are given V, T, and P, but none have the correct units for our value of R, so we convert each into the proper units.

$$\mathbf{V = 5.67\ mL \times \frac{1\ L}{1000\ mL} = 0.00567\ L} \qquad \mathbf{T = 26\ ^oC + 273 = 299\ K}$$

$$\mathbf{P = 782\ mm\ Hg \times \frac{1.00\ atm}{760\ mm\ Hg} = 1.03\ atm}$$

Next, rearrange the ideal gas law to solve for n.

$$\mathbf{n = \frac{PV}{RT} = \frac{(1.03\ atm)\ (\ 0.00567\ L)}{(0.0821\ L \cdot atm \cdot K^{-1} \cdot mol^{-1})\ (299\ K)} = 2.38 \times 10^{-4}\ mol\ CO_2}$$

To find grams, apply the molar mass (M_m = 44.01 g/mol) as a conversion factor as done in Appendix A,

$$\mathbf{2.38 \times 10^{-4}\ mol\ CO_2 \times \frac{44.01\ g\ CO_2}{mol\ CO_2} = 0.0105\ g\ CO_2}$$

Comment:

Notice that five and two-thirds mL of a gas seems like a very small amount in terms of mass! That is because we are used to weighing out solids and liquids, which are much denser than gases.

B.4 EXERCISES

1. What volume does 0.50 moles of CO_2 occupy at 725 mm Hg and 25 ºC?

2. How many moles of He occupy a 2.50-L flask whose pressure is 945 mm Hg at 75 ºC?

3. What is the pressure exerted by 28.8 g of N_2 contained in a 4.25 L-flask at 0 ºC?

4. What volume does 5.8 moles of O_2 occupy at 285 mm Hg and -78 ºC?

5. What is the temperature of 5.0 moles of N_2 contained in a 20.0 L-tank at a pressure of 7.5 atm?

6. What volume does 6.32 g of NH_3 occupy at 745 mm Hg and 25 ºC?

7. How many moles of CH_4 occupy a 10.0-L tank whose pressure is 3.5 atm at 30 ºC?

8. What volume does 0.45 g of Ar occupy at 1.25 atm and 27 ºC?

9. What is the pressure exerted by 3.5 moles of H_2 contained in a 2.0-L tank at 27 ºC?

10. What volume does 0.75 moles of N_2 occupy at 760 mm Hg and 0 ºC?

11. What is the temperature of 7.65 g of He contained in a 6.25 L flask at a pressure of 1.75 atm?

12. How many moles of HCl gas occupy a 4.5 L tank whose pressure is 1875 mm Hg at 27 ºC?

13. For this question, note that M_m = g/mol and density, d = mass/volume.

a) What is the density of helium in g/L at 1.00 atm and 27 ºC?

b) What is the density of nitrogen in g/L at 1.00 atm and 27 ºC?

ANSWERS:

1. 13 L
2. 0.109 mol
3. 5.42 atm
4. 2.5×10^2 L
5. 365 K = 92 ^oC
6. 9.26 L
7. 1.4 mol
8. 0.22 L
9. 43 atm
10. 17 L
11. 69.7 K = -203 ^oC
12. 0.45 mol
13. a) 0.163 g/L
13. b) 1.14 g/L