5. THE GASEOUS STATE

■ Solutions to Exercises

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.

5.1   First, the conversion to atm (57 kPa = 57 x 10^3 Pa).

$$57 \times 10^3 \text{ Pa } \times \frac{1 \text{ atm}}{1.01325 \times 10^5 \text{ Pa}} = 0.5\underline{6}2 = 0.56 \text{ atm}$$

Next, convert to mmHg.

$$57 \times 10^3 \text{ Pa } \times \frac{760 \text{ mmHg}}{1.01325 \times 10^5 \text{ Pa}} = 4\underline{2}7.5 = 4.3 \times 10^2 \text{ mmHg}$$

5.2   Application of Boyle's law gives

$$V_f = V_i \times \frac{P_i}{P_f} = 20.0 \text{ L } \times \frac{1.00 \text{ atm}}{0.830 \text{ atm}} = 24.\underline{0}96 = 24.1 \text{ L}$$

5.3   First, convert the temperatures to the Kelvin scale.

$T_i = (19 + 273) = 292 \text{ K}$

$T_f = (25 + 273) = 298 \text{ K}$

Following is the data table

$V_i = 4.38 \text{ dm}^3$	$P_i = 101 \text{ kPa}$	$T_i = 292 \text{ K}$
$V_f = ?$	$P_f = 101 \text{ kPa}$	$T_f = 298 \text{ K}$

Apply Charles's law to obtain

$$V_f = V_i \times \frac{T_i}{T_f} = 4.38 \text{ dm}^3 \times \frac{298 \text{ K}}{292 \text{ K}} = 4.4\underline{7}0 = 4.47 \text{ dm}^3$$

5.4   First, convert the temperatures to kelvins.

$T_i = (24 + 273) = 297 \text{ K}$

$T_f = (35 + 273) = 308 \text{ K}$

Following is the data table

$V_i = 5.41 \text{ dm}^3$	$P_i = 101.5 \text{ kPa}$	$T_i = 297 \text{ K}$
$V_f = ?$	$P_f = 102.8 \text{ kPa}$	$T_f = 308 \text{ K}$

Apply both Boyle's law and Charles's law combined to get

$$V_f = V_i \times \frac{P_i}{P_f} \times \frac{T_f}{T_i} = 5.41 \text{ dm}^3 \times \frac{101.5 \text{ kPa}}{102.8 \text{ kPa}} \times \frac{308 \text{ K}}{297 \text{ K}} = 5.5\underline{3}9 = 5.54 \text{ dm}^3$$

5.5   Use the ideal gas law, PV = nRT, and solve for n:

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

Note that everything in parentheses is constant. Therefore, you can write

$n = \text{constant} \times P$

Or, expressing this as a proportion, you get

$n \propto P$

5.6   First, convert the mass of O_2 to moles O_2 (molar mass 32.00 g/mol) and convert temperature to kelvins.

$$T = 23 + 273 = 296 \text{ K}$$

$$3.03 \text{ kg } O_2 \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} = 94.\underline{6}88 \text{ mol } O_2$$

Summarize the data in a table.

Variable	Value
P	?
V	50.0 L
T	296 K
n	94.688 mol

Solve the ideal gas equation for P, and substitute the data to get

$$P = \frac{nRT}{V} = \frac{(94.6\underline{8}8)(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(296 \text{ K})}{50.0 \text{ L}} = 46.\underline{0}0 = 46.0 \text{ atm}$$

5.7   The data given are

Variable	Value
P	752 mmHg $\times \dfrac{1 \text{ atm}}{760 \text{ mmHg}} = 0.98\underline{9}47$ atm
V	1 L (exact number)
T	(21 + 273) = 294 K
n	?

Using the ideal gas law, solve for n, the moles of helium.

$$n = \frac{PV}{RT} = \frac{(0.98\underline{9}47 \text{ atm})(1 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(294 \text{ K})} = 0.041\underline{0}1 \text{ mol}$$

Now convert mol He to grams

$$0.041\underline{0}1 \text{ mol He } \times \frac{4.00 \text{ g He}}{1 \text{ mol He}} = 0.16\underline{4}04 \text{ g He}$$

Therefore, the density of He at 21°C and 752 mmHg is 0.164 g/L.

The difference in mass between one liter of air and one liter of helium is

Mass air - mass He = 1.188 g - 0.16404 g = 1.02396 = 1.024 g difference

5.8   Tabulate the values of the variables:

Variable	Value
P	0.862 atm
V	1 L (exact number)
T	(25 + 273) = 298 K
n	?

From the ideal gas law, PV = nRT, you obtain

$$n = \frac{PV}{RT} = \frac{(0.862 \text{ atm})(1 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(298 \text{ K})} = 0.0352\underline{5} \text{ mol}$$

Dividing the mass of the vapor by moles gives you the mass per mole (the molar mass).

$$\text{Molar mass} = \frac{\text{grams vapor}}{\text{moles vapor}} = \frac{2.26 \text{ g}}{0.03525 \text{ mol}} = 64.\underline{1}14 \text{ g/mol}$$

Therefore, the molecular weight is 64.1 amu.

5.9   First, determine the number of moles of Cl_2 from the mass of HCl (molar mass 36.46 g/mol) and from the stoichiometry of the chemical equation:

$$9.41 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \times \frac{5 \text{ mol Cl}_2}{16 \text{ mol HCl}} = 0.0806\underline{5}3 \text{ mol Cl}_2$$

Tabulate the values of the variables:

Variable	Value
P	787 mmHg $\times \dfrac{1 \text{ atm}}{760 \text{ mmHg}}$ = 1.0355 atm
T	(40 + 273) K = 313 K
n	0.0806\underline{5}3 mol
V	?

Rearrange the ideal gas law to obtain V:

$$V = \frac{nRT}{P} = \frac{(0.080653 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(313 \text{ K})}{1.0355 \text{ atm}} = 2.0\underline{0}1 = 2.00 \text{ L}$$

5.10   Each gas obeys the ideal gas law. In each case, convert grams to moles and substitute
into the ideal gas law to determine the partial pressure of each.

$$1.031 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} = 0.0322\underline{1}88 \text{ mol } O_2$$

$$P = \frac{nRT}{V} = \frac{(0.0322188)(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(291 \text{ K})}{10.0 \text{ L}} = 0.076\underline{9}36 \text{ atm}$$

$$0.572 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} = 0.012\underline{9}97 \text{ mol } CO_2$$

$$P = \frac{nRT}{V} = \frac{(0.012997)(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(291 \text{ K})}{10.0 \text{ L}} = 0.031\underline{0}36 \text{ atm}$$

The total pressure is equal to the sum of the partial pressures:

$$P = P_{O_2} + P_{CO_2} = 0.076\underline{9}36 + 0.031\underline{0}36 = 0.107\underline{9}7 = 0.1080 \text{ atm}$$

The mole fraction of oxygen in the mixture is

$$\text{Mole fraction } O_2 = \frac{P_{N_2}}{P} = \frac{0.07694 \text{ atm}}{0.108\underline{0}2 \text{ atm}} = 0.71\underline{2}2 = 0.712$$

5.11   Determine the number of moles of O_2 from the mass of $KClO_3$ and from the
stoichiometry of the chemical reaction.

$$1.300 \text{ g } KClO_3 \times \frac{1 \text{ mol } KClO_3}{122.5 \text{ g } KClO_3} \times \frac{3 \text{ mol } O_2}{2 \text{ mol } KClO_3} = 0.0159\underline{1}84 \text{ mol } O_2$$

The vapor pressure of water at 23°C is 21.1 mmHg (Table 5.6). Find the partial
pressure of O_2 using Dalton's law:

$$P = P_{O_2} + P_{H_2O}$$

$$P_{O_2} = P - P_{H_2O} = (745 - 21.1) \text{ mmHg} = 72\underline{3}.9 \text{ mmHg}$$

Solve for the volume using the ideal gas law.

Variable	Value

P 723.9 mmHg $\times$ $\dfrac{1 \text{ atm}}{760 \text{ mmHg}}$ = 0.95$\underline{2}$5 atm

V ?
T (23 + 273) = 296 K
n 0.0159184

From the ideal gas law, PV = nRT, you have

$$V = \frac{nRT}{P} = \frac{(0.0159184 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(296 \text{ K})}{0.9525 \text{ atm}}$$

$$= 0.40\underline{5}9 = 0.406 \text{ L}$$

5.12   The absolute temperature is (22 + 273) = 293 K. In SI units, the molar mass of carbon tetrachloride, CCl_4, is 153.8 x 10^{-3} kg/mol. Therefore,

$$u = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3 \times 8.31 \text{ kg} \cdot \text{m}^2/(\text{s}^2 \cdot \text{K} \cdot \text{mol}) \times 295 \text{ K}}{153.8 \times 10^{-3} \text{ kg/mol}}}$$

$$= 21\underline{8}.7 = 219 \text{ m/s}$$

5.13   Determine the rms molecular speed for N_2 at 455°C (728 K):

$$u = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3 \times 8.31 \text{kg} \cdot \text{m}^2/(\text{s}^2 \cdot \text{K} \cdot \text{mol}) \times 728 \text{ K}}{28.02 \times 10^{-3} \text{ kg/mol}}} = 80\underline{4}.81 \text{m/s}$$

After writing this equation with the same speed for H_2, square both sides and solve for T. The molar mass of H_2 in SI units is 2.016 x 10^{-3} kg/mol. Therefore,

$$T = \frac{u^2 M}{3R} = \frac{(804.81)^2 (2.016 \times 10^{-3} \text{ kg/mol})}{(3)(8.31 \text{kg} \cdot \text{m}^2/\text{s}^2 \cdot \text{K} \cdot \text{mol})} = 52.\underline{3}79 = 52.4 \text{ K}$$

Because the average kinetic energy of a molecule is proportional to only T, the temperature at which an H_2 molecule has the same average kinetic energy as an N_2 molecule at 455°C is exactly the same temperature, 455°C.

5.14 The two rates of effusion are inversely proportional to the square roots of their molar masses, so you can write

$$\frac{\text{Rate of effusion of } O_2}{\text{Rate of effusion of He}} = \sqrt{\frac{M_m(\text{He})}{M_m(O_2)}}$$

where $M_m(\text{He})$ is the molar mass of He (4.00 g/mol) and $M_m(O_2)$ is the molar mass of O_2 (32.00 g/mol). Substituting these values into the formula gives

$$\frac{\text{Rate of effusion of } O_2}{\text{Rate of effusion of He}} = \sqrt{\frac{4.00 \text{ g/mol}}{32.00 \text{ g/mol}}} = 0.35\underline{3}55$$

Rearranging gives

rate of effusion of O_2 = 0.35$\underline{3}$55 x rate of effusion of He.

Now, the problem states that the rate of effusion can be given in terms of volume of gas effused per second, so

$$\frac{\text{Volume of } O_2}{\text{time for } O_2} = 0.35\underline{3}55 \text{ x } \frac{\text{Volume of He}}{\text{time for He}}$$

Substituting in the values gives

$$\frac{10.0 \text{ mL}}{\text{time for } O_2} = 0.35\underline{3}55 \text{ x } \frac{10.0 \text{ mL}}{3.52 \text{ s}}$$

Rearranging gives

$$\text{time for } O_2 = \frac{3.52 \text{ s}}{0.35\underline{3}55} = 9.9\underline{5}6 = 9.96 \text{ s}$$

5.15 The problem states the rate of effusion is inversely proportional to the time it takes for a gas to effuse, so you can write

$$\frac{\text{Rate of effusion of } H_2}{\text{Rate of effusion of gas}} = \frac{\text{time for gas}}{\text{time for } H_2} = \sqrt{\frac{M_m(\text{gas})}{M_m(H_2)}} = 4.67$$

Rearranging and solving for $M_m(H_2)$ gives

$$M_m(\text{gas}) = (4.67)^2 \text{ x } M_m(H_2) = (4.67)^2 \text{ x } 2.016 \text{ g/mol} = 43.\underline{9}6 = 44.0 \text{ g/mol}$$

Thus, the molecular weight of the gas is 44.0 amu.

5.16   From Table 5.7, a = 5.570 L^2•atm/mol^2 and b = 0.06499 L/mol. Substitute these values into the van der Waals equation along with R = 0.08206 L•atm/K•mol, T = 273.2 K, n = 1.000 mol, and V = 22.41 L.

$$P = \frac{nRT}{(V - nb)} - \frac{n^2a}{V^2}$$

$$P = \frac{(1.00 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(273.2 \text{ K})}{22.41 \text{L} - (1.00 \text{ mol})(0.06499 \text{ L/mol})} -$$

$$\frac{(1.00 \text{ mol})^2 (5.570 \text{ L}^2 \cdot \text{atm/mol}^2)}{(22.41 \text{L})^2}$$

$$= 1.00\underline{33} - 0.0110\underline{91} = 0.99\underline{221} = 0.992 \text{ atm}$$

Using the ideal gas law, P = 1.00$\underline{04}$ atm (larger).

■ Answers to Concept Checks

5.1   The general relationship between pressure (P) and the height (h) of a liquid column in a barometer is P = gdh, where g is the constant acceleration of gravity and d is the density. Examination of the relationship indicates that, for a given pressure, as the density of the liquid in the barometer decreases, the height of the liquid must increase. In order to make this relationship more apparent, you can rearrange the equation to:

$$gh = \frac{P}{d}$$

Keeping in mind that you are conducting the experiment at constant pressure and that gravity is a constant, this mathematical relationship demonstrates that the height of the liquid in the barometer is inversely proportional to the density of the liquid in the barometer.

$$(h \propto \frac{1}{d})$$

This inverse relationship means that, as the height of the liquid decreases, the density of the liquid must increase. Since the density of mercury is greater than the density of water, the barometer with the water will have the higher column.

5.2 a. In the first step, when the temperature decreases, the pressure will also decrease. This is because, according to the combined gas law, the pressure is directly proportional to the temperature ($P \propto T$). In the second step, when the volume increases, the pressure will decrease, since according to Boyles law, pressure and volume are inversely related ($P \propto 1/V$). Both changes result in the pressure decreasing, so the final pressure will be less than the starting pressure.

   b. In the first step, when the temperature increases, the pressure will also increase. This is because, according to the combined gas law, pressure is directly proportional to the temperature ($P \propto T$). In the second step, when the volume decreases, the pressure will increase, since according to the ideal gas law, pressure and volume are inversely related ($P \propto 1/V$). Both changes result in the pressure increasing, so the final pressure will be greater than the starting pressure.

5.3 a. According to Avogadro's law, equal volumes of any two (or more) gases at the same temperature and pressure contain the same number of molecules (or atoms in this case). Therefore, all three flasks contain the same number of atoms.

   b. Since density is mass divided by volume, and all three flasks have the same volume (3.0 L), the gas with the largest molar mass, xenon (Xe), will have the greatest density.

   c. According to the ideal gas law, $PV = nRT$, pressure is directly proportional to the temperature. Since the helium flask is being heated, it will have the highest pressure.

   d. Since the three flasks started with the same number of atoms, and hence the same number of moles, they would all still have the same number of moles no matter how the temperatures of the flasks were changed.

5.4 a. In a mixture of gases, each gas exerts the pressure it would exert if it were the only gas in the flask. The pressure of H_2 is the same whether it is in the flask by itself or with the Ar. Therefore, the pressure of H_2 does not change.

   b. According to the ideal gas law, $PV = nRT$, pressure (P) is directly proportional to the number of moles (n). Since the number of moles of H_2 and the number of moles of Ar are equal, their pressures are also equal.

   c. The total pressure is equal to the sum of the pressures of the H_2 gas and the Ar gas in the container. The total pressure will also be equal to twice the pressure of the H_2 gas when it was in the flask by itself. It is also equal to twice the pressure that the Ar gas would exert if it were in the flask by itself.

5.5 a. The rate of effusion is inversely proportional to the square root of the molecular weight of the gas at constant temperature and pressure. Thus, He (molecular weight 4.00 amu) will diffuse faster than Ar (molecular weight 39.95 amu) and reach the end of the tube first.

b. The speed of an atom is directly proportional to the absolute temperature. If you raise the temperature of the Ar, you can make it reach the end of the tube at the same time as the He.

5.6 a. If the real gas molecules occupy a relatively small volume, then the volume of the gas is essentially equal to the volume of the container, the same as for an ideal gas. However, if there were large intermolecular attractions, the pressure would be less than for an ideal gas. Therefore, the pressure would be greater for the ideal gas.

b. If the real gas molecules occupy a relatively large volume, then the volume available for the gas is less than for an ideal gas, and the pressure would be greater. If there are negligible intermolecular attractions, then the pressure is essentially the same as for an ideal gas. Overall, the pressure would be less for the ideal gas.

c. Since the effect of molecular volume and intermolecular attractions on the pressure of a real gas are opposite, you cannot determine how the pressure of the two gases compare.

■ Answers to Review Questions

5.1 Pressure is the force exerted per unit area of surface. Force is further defined as mass multiplied by acceleration. The SI unit of mass is kg, of acceleration is m/s^2, and of area is m^2. Therefore, the SI unit of pressure (Pascal) is given by

$$\text{Pressure} = \frac{\text{force}}{\text{area}} = \frac{\text{mass x acceleration}}{\text{area}} = \frac{\text{kg x m/s}^2}{m^2} = \frac{\text{kg}}{m \cdot s^2} = \text{Pa}$$

5.2 A manometer is a device that measures the pressure of a gas in a vessel. The gas pressure in the flask is proportional to the difference in heights between the liquid levels in the manometer (Figure 5.4).

5.3 The general relationship between the pressure (P) and the height (h) of the liquid in a manometer is P = dgh. Therefore, the variables that determine the height of the liquid in a manometer are the density (d) of the liquid and the pressure of the gas being measured. The acceleration of gravity (g) is a constant, 9.81 m/s^2.

5.4 From Boyle's law, PV = constant. Because this is true for conditions P_i and V_i as well as conditions P_f and V_f, we can write

$P_f V_f = P_i V_i = \text{constant}$

Dividing both sides of this equation by P_f gives

$$V_f = V_i \times \frac{P_i}{P_f}$$

5.5   A linear relationship between variables such as x and y is given by the mathematical relation

    $$y = a + bx$$

The variable y is directly proportional to x only if a = 0.

5.6   First, find the equivalent of absolute zero on the Fahrenheit scale. Converting -273.15°C to degrees Fahrenheit, you obtain -459.67°F. Since the volume of a gas varies linearly with the temperature, you get the following linear relationship.

    $$V = a + bt_F$$

where t_F is the temperature on the Fahrenheit scale. Since the volume of a gas is zero at absolute zero, you get

    $$0 = a + b(-459.67), \text{ or } a = 459.67b$$

The equation can now be rewritten as

    $$V = 459.67b + bt = b(459.67 + t_F) = bT_F$$

where T_F is the temperature in the new absolute scale based on the Fahrenheit scale. The relationship is

    $$T_F = t_F + 459.67$$

5.7   From Charles's law, V = constant x T. Because this is true for conditions T_i and V_i as well as conditions T_f and V_f, we can write

    $$\frac{V_f}{T_f} = \frac{V_i}{T_i} = \text{constant}$$

Multiplying both sides of the equation by T_f gives

    $$V_f = V_i \times \frac{T_f}{T_i}$$

5.8   Avogadro's law states that equal volumes of any two gases at the same temperature and pressure contain the same number of molecules. The law of combining volumes states that the volumes of reactant gases at a given pressure and temperature are in ratios of small whole numbers. The combining-volume law may be explained from Avogadro's law using the reaction $N_2 + 3H_2 \rightarrow 2NH_3$ as follows: In Avogadro's terms, this equation says that Avogadro's number of N_2 molecules reacts with three times Avogadro's number of H_2 molecules to form two times Avogadro's number of NH_3 molecules. From Avogadro's law, it follows that one volume of N_2 reacts with three volumes of H_2 to form two volumes of NH_3. This result is true for all gas reactions.

5.9   The standard conditions are 0°C and one atm pressure (STP).

5.10   The molar gas volume, V_m, is the volume of one mole of gas at any given temperature and pressure. At standard conditions (STP), the molar gas volume equals 22.4 L.

5.11   Boyle's law ($V \propto 1/P$) and Charles's law ($V \propto T$) can be combined and expressed in a single statement: The volume occupied by a given amount of gas is proportional to the absolute temperature divided by the pressure. In equation form, this is

$$V = \text{constant} \times \frac{T}{P}$$

The constant is independent of temperature and pressure, but does depend on the amount of gas. For one mole, the constant will have a specific value, denoted as R. The molar volume, V_m, is

$$V_m = R \times \frac{T}{P}$$

Because V_m has the same value for all gases, we can write this equation for n moles of gas if we multiply both sides by n. This yields the equation

$$nV_m = \frac{nRT}{P}$$

Because V_m is the volume per mole, nV_m is the total volume V. Substituting gives

$$V = \frac{nRT}{P} \qquad \text{or} \qquad PV = nRT$$

5.12   The variables in the ideal gas law are P, V, n, and T. The SI units of these variables are pascals (P), cubic meters (V), moles (n), and kelvins (T).

5.13   Use the value of R from Table 5.5, and the conversion factor 1 atm = 760 mmHg. This gives

$$0.082058 \frac{L \cdot atm}{K \cdot mol} \times \frac{760 \, mmHg}{1 \, atm} = 62.36\underline{4}0 = 62.364 \frac{L \cdot mmHg}{K \cdot mol}$$

5.14   Six empirical gas laws can be obtained. They can be stated as follows:

   $P \times V = $ constant   (T and n constant)

   $\dfrac{P}{T} = $ constant   (V and n constant)

   $\dfrac{P}{n} = $ constant   (T and V constant)

   $\dfrac{V}{T} = $ constant   (P and n constant)

   $\dfrac{V}{n} = $ constant   (P and T constant)

   $n \times T = $ constant   (P and V constant)

5.15   The postulates and supporting evidence are the following:

   (1) Gases are composed of molecules whose sizes are negligible compared with the distance between them.

   (2) Molecules move randomly in straight lines in all directions and at various speeds.

   (3) The forces of attraction or repulsion between two molecules (intermolecular forces) in a gas are very weak or negligible, except when they collide.

   (4) When molecules collide with one another, the collisions are elastic.

   (5) The average kinetic energy of a molecule is proportional to the absolute temperature.

   One example of evidence that supports the kinetic theory of gases is Boyle's law. Constant temperature means the average molecular force from collision remains constant. If you increase the volume, you decrease the number of collisions per unit wall area, thus lowering the pressure in accordance with Boyle's law. Another example is Charles's law. If you raise the temperature, you increase the average molecular force from a collision with the wall, thus increasing the pressure. For the pressure to remain constant, it is necessary for the volume to increase so the frequency of collisions with the wall decreases. Thus, when you raise the temperature of a gas while keeping the pressure constant, the volume increases in accordance with Charles's law.

5.16   Boyle's law requires the temperature be constant. Postulate 5 of the kinetic theory holds
        that the average kinetic energy of a molecule is constant at constant temperature.
        Therefore, the average molecular force from collisions is constant. If we increase the
        volume of a gas, this decreases the number of molecules per unit volume and so
        decreases the frequency of collisions per unit wall area, causing the pressure to
        decrease in accordance with Boyle's law.

5.17   According to kinetic theory, the pressure of a gas results from the bombardment of
        container walls by molecules.

5.18   The rms speed of a molecule equals $(3RT/M_m)^{1/2}$, where M_m is the molar mass of the
        gas. The rms speed does not depend on the molar volume.

5.19   A gas appears to diffuse more slowly because it never travels very far in one direction
        before it collides with another molecule and moves in another direction. Thus, it must
        travel a very long, crooked path as the result of collisions.

5.20   Effusion is the process in which a gas flows through a small hole in a container. It
        results from the gas molecules encountering the hole by chance, rather than by
        colliding with the walls of the container. The faster the molecules move, the more likely
        they are to encounter the hole. Thus, the rate of effusion depends on the average
        molecular speed, which depends inversely on molecular mass.

5.21   The behavior of a gas begins to deviate significantly from that predicted by the ideal
        gas law at high pressures and relatively low temperatures.

5.22   The constant "a" is the proportionality constant in the van der Waals equation related
        to intermolecular forces The term "nb" represents the volume occupied by n moles of
        molecules.

■ Answers to Conceptual Problems

5.23 a. The volume of the tire and the amount of air in the tire remain constant. From the ideal
         gas law, PV = nRT, under these conditions the pressure will vary directly with the
         temperature ($P \propto T$). Thus, on a cold day, you would expect the pressure in the tires to
         decrease, and they would appear flatter.

     b. Aerosol cans are filled with a fixed amount of gas in a constant volume. From the ideal
         gas law, under these conditions the pressure will vary directly with the temperature
         ($P \propto T$). If you put an aerosol can in a fire, you will increase the temperature, and thus
         the pressure. If the pressure gets high enough, the can will explode.

     c. As the water bottle sits in the sun, the liquid water warms up. As the temperature of the
         water increases, so does its vapor pressure (Table 5.6). If the pressure gets high
         enough, it will pop the lid off the bottle.

    d. The amount of air in the balloon and the temperature remain constant. From the ideal gas law under these conditions, the pressure is inversely proportional to the volume (P ∝ 1/V). Thus, as you squeeze the balloon, you decrease the volume, resulting in an increase in pressure. If you squeeze hard enough and make the volume small enough, the balloon will pop.

5.24   Since each of the flasks is identical and each contains an equal amount of gas, the initial pressures in the N_2 flask, the O_2 flask, and the He flask will be the same. After the changes, the pressure in the He flask would be highest, with a pressure equal to three times the original. Next would be the O_2 flask, with a pressure equal to two times the original. Last would be the N_2 flask, with a pressure equal to one-half the original.

5.25 a. The pressure and volume of a gas are inversely proportional; therefore, an increase by a factor of two in pressure would decrease the volume by ½ (C to D).

    b. The pressure and volume of a gas are inversely proportional; therefore, a decrease by a factor of two in pressure would double the volume (C to A).

    c. The volume and temperature of a gas are directly proportional; therefore an increase in kelvin temperature by a factor of 1.5 would result in an increase in volume by a factor of 1.5 (C to B).

    d. Since the piston can move, the pressure would not change (it would be equal to the starting pressure). The volume of gas is directly proportional to the number of moles; therefore an increase in the number of moles by a factor of two would cause the volume increase by two (C to A).

5.26 a. Since 1.0 out of the 3.0 moles of gas in the container is N_2, the fraction of the pressure due to N_2 is 1/3.

    b. Mole fractions are not a function of temperature, so nothing would happen.

    c. You would expect the pressure to be higher for two reasons. First, the water would occupy some volume, reducing the volume available for the gas to occupy. Thus, according to Boyle's law, as volume decreases, pressure increases. Second, after a time, the water would evaporate, and the vapor pressure due to the water would contribute to the total pressure, thereby increasing it.

    d. Yes, there is enough information in the problem to calculate the pressure in the flask, but you would also need to know the vapor pressure of water at 22.0°C (Table 5.6).

5.27 a. The container with the O_2 has the greater density, since the molar mass of O_2 (32.00 g/mol) is greater than that of H_2 (2.016 g/mol).

    b. Since the H_2 molecules are lighter, they will be moving faster.

    c. Both containers have the same number of molecules (Avogadro's law).

d. The pressure in each of the containers will not change when the valve is opened. Each container starts with the same pressure. Since the total volume remains constant, the pressure will not change.

e. The fraction of the total pressure due to the H_2 would now be 1/4.

5.28 a. Pressure decreases as you increase in altitude. Thus, the pressure at 6000 m is less than the pressure at 3000 m. For two identical balloons, the balloon at 6000 m will have the greater volume (Boyle's law).

b. In order to calculate the volume of each balloon, you would need the temperature and pressure on the ground and the temperature and pressure at their respective heights.

5.29 In order to double the volume you could reduce the pressure by 1/2. You could also increase the temperature, but you cannot determine the final temperature without knowing the initial temperature.

5.30 a. If you assume that the flasks are at ordinary room temperature, say 25°C, then there would be approximately one mole of He (at 1.0 atm) and 1.5 mol of Xe (at 1.5 atm). Thus, the F_2 flask (with 2.0 mol) would contain the greatest number of moles of gas.

b. You could either decrease the volume, increase the temperature, or both.

■ Solutions to Practice Problems

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.

5.31 Use the conversion factor 1 atm = 760 mmHg.

$$0.047 \text{ atm} \times \frac{760 \text{ mmHg}}{1 \text{ atm}} = 3\underline{5}.7 = 36 \text{ mmHg}$$

5.32 Use the conversion factors 1 atm = 760 mmHg and 1 atm = 101325 Pa, or 101.325 kPa

$$259 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} \times \frac{101.325 \text{ kPa}}{1 \text{ atm}} = 34.\underline{5}3 = 34.5 \text{ kPa}$$

5.33   Using Boyle's law, solve for V_f of the neon gas at 1.292 atm pressure.

$$V_f = V_i \times \frac{P_i}{P_f} = 3.15 \text{ L} \times \frac{0.951 \text{ atm}}{1.292 \text{ atm}} = 2.3\underline{1}8 = 2.32 \text{ L}$$

5.34   Using Boyle's law, solve for V_f of the helium gas at 632 mmHg.

$$V_f = V_i \times \frac{P_i}{P_f} = 2.68 \text{ L} \times \frac{789 \text{ mmHg}}{632 \text{ mmHg}} = 3.3\underline{4}57 = 3.35 \text{ L}$$

5.35   Using Boyle's law, let V_f = volume at 0.974 atm (P_f), V_i = 50.0 L, and P_i = 19.8 atm.

$$V_f = V_i \times \frac{P_i}{P_f} = 50.0 \text{ L} \times \frac{19.8 \text{ atm}}{0.974 \text{ atm}} = 10\underline{1}6.4 = 1.02 \times 10^3 \text{ L}$$

5.36   Using Boyle's law, let V_f = volume at 1.584 atm (P_f), V_i = 8.58 m^3, and P_i = 1.020 atm.

$$V_f = V_i \times \frac{P_i}{P_f} = 8.58 \text{ m}^3 \times \frac{1.020 \text{ atm}}{1.584 \text{ atm}} = 5.5\underline{2}5 = 5.53 \text{ L}$$

5.37   Using Boyle's law, let P_i = pressure of 315 cm^3 of gas, and solve for it:

$$P_i = P_f \times \frac{V_f}{V_i} = 2.51 \text{ kPa} \times \frac{0.0457 \text{ cm}^3}{315 \text{ cm}^3} = 3.6\underline{4}1 \times 10^{-4} = 3.64 \times 10^{-4} \text{ kPa}$$

5.38   Using Boyle's law, let P_f = final pressure of 27.0 dm^3 of gas, and solve for it:

$$P_f = P_i \times \frac{V_i}{V_f} = 101 \text{ kPa} \times \frac{456 \text{ dm}^3}{27.0 \text{ dm}^3} = 17\underline{0}5.7 = 1.71 \times 10^3 \text{ kPa}$$

5.39   Use Charles's law: T_i = 18°C + 273 = 291 K, and T_f = 0°C + 273 = 273 K.

$$V_f = V_i \times \frac{T_f}{T_i} = 3.92 \text{ mL} \times \frac{273 \text{ K}}{291 \text{ K}} = 3.6\underline{7}7 = 3.68 \text{ mL}$$

5.40   Use Charles's law: T_i = 0°C + 273 = 273 K, and T_f = 20°C + 273 = 293 K.

$$V_f = V_i \times \frac{T_f}{T_i} = 22.41 \text{ L} \times \frac{293 \text{ K}}{273 \text{ K}} = 24.\underline{0}51 = 24.1 \text{ L}$$

5.41   Use Charles's law: $T_i = 22°C + 273 = 295$ K, and $T_f = -197°C + 273 = 76$ K.

$$V_f = V_i \times \frac{T_f}{T_i} = 2.54 \text{ L} \times \frac{76 \text{ K}}{295 \text{ K}} = 0.6\underline{5}4 = 0.65 \text{ L}$$

5.42   Use Charles's law: $T_i = 0°C + 273 = 273$ K, and $T_f = 25°C + 273 = 298$ K.

$$V_f = V_i \times \frac{T_f}{T_i} = 4.83 \text{ L} \times \frac{298 \text{ K}}{273 \text{ K}} = 5.2\underline{7}2 = 5.27 \text{ L}$$

5.43   Use Charles's law: $T_i = 25°C + 273 = 298 \text{ K}_i$, and V_f is the difference between the vessel's volume of 39.5 cm^3 and the 7.7 cm^3 of ethanol that is forced into the vessel.

$$T_f = T_i \times \frac{V_f}{V_i} = \frac{(39.5 - 7.7) \text{ cm}^3}{39.5 \text{ cm}^3} = 23\underline{9}.9 \text{ K } (-3\underline{3}.2 \text{ or } -33°C)$$

5.44   Use Charles's law: $T_i = 18°C + 273 = 291$ K, $V_i = 62.3$ cm^3, and $V_f = 50.9$ cm^3.

$$T_f = T_i \times \frac{V_f}{V_i} = 291 \text{ K} \times \frac{50.9 \text{ cm}^3}{62.3 \text{ cm}^3} = 23\underline{7}.7 = 238 \text{ K } (-3\underline{5}.3 \text{ or } -35°C)$$

5.45   Use the combined law: $T_i = 31°C + 273 = 304$ K, and $T_f = 0°C + 273 = 273$ K.

$$V_f = V_i = \times \frac{P_i}{P_f} \times \frac{T_f}{T_i} = 35.5 \text{ mL} \times \frac{753 \text{ mmHg}}{760 \text{ mmHg}} \times \frac{273 \text{ K}}{304 \text{ K}} = 31.\underline{5}8 = 31.6 \text{ mL}$$

5.46   Use the combined law: $T_i = 23°C + 273 = 296$ K, and $T_f = 0°C + 273 = 273$ K.

$$V_f = V_i = \times \frac{P_i}{P_f} \times \frac{T_f}{T_i} = 3.84 \text{ mL} \times \frac{785 \text{ mmHg}}{760 \text{ mmHg}} \times \frac{273 \text{ K}}{296 \text{ K}} = 3.6\underline{5}8 = 3.66 \text{ mL}$$

5.47   The balanced equation is

$$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$$
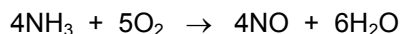
The ratio of moles of NH_3 to moles of NO = 4 to 4, or 1 to 1, so one volume of NH_3 will produce one volume of NO at the same temperature and pressure.

5.48   The balanced equation is

$$CO_2 + 3H_2 \rightarrow CH_3OH + H_2O$$
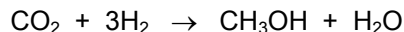
The ratio of moles of H_2 to moles of CO_2 = 3 to 1, so three volumes of H_2 are required to react with one volume of CO_2 at the same temperature and pressure.

5.49   Solve the ideal gas law for V:

$$V = \frac{nRT}{P} = nRT\left(\frac{1}{P}\right)$$

If the temperature and number of moles are held constant, then the product nRT is constant, and volume is inversely proportional to pressure:

$$V = \text{constant} \times \frac{1}{P}$$

5.50   Solve the ideal gas law for V:

$$V = \frac{nR}{P} \times T$$

If n and P are held constant, the nR/P quotient is a constant. The equation may now be written

$$V = \text{constant} \times T$$

5.51   Calculate the moles of oxygen, and then solve the ideal gas law for P:

$$n = 91.3 \text{ g} \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 2.8\underline{5}3 \text{ mol O}_2$$

$$P = \frac{nRT}{V} = \frac{(2.8\underline{5}3 \text{ mol})(0.08206 \text{ L} \cdot \text{atm}/\text{K} \cdot \text{mol})(294 \text{ K})}{8.58 \text{ L}} = 8.0\underline{2}2 = 8.02 \text{ atm}$$

5.52   Calculate the moles of methane, and then solve the ideal gas law for P:

$$n = 7.13 \text{ g} \times \frac{1 \text{ mol CH}_4}{16.00 \text{ g CH}_4} = 0.44\underline{5}6 \text{ mol}$$

$$P = \frac{nRT}{V} = \frac{(0.4456 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(292 \text{ K})}{5.00 \text{ L}}$$

$$= 2.1\underline{3}55 = 2.14 \text{ atm}$$

5.53   Using the moles of chlorine, solve the ideal gas law for V:

$$V = \frac{nRT}{P} = \frac{(3.50 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(307 \text{ K})}{4.00 \text{ atm}} = 22.\underline{0}4 = 22.0 \text{ L}$$

5.54   Calculate the moles of oxygen, and then solve the ideal gas law for V:

$$n = 5.67 \text{ g} \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 0.17\underline{7}18 \text{ mol}$$

$$V = \frac{nRT}{P} = \frac{(0.17718 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(296 \text{ K})}{0.985 \text{ atm}}$$

$$= 4.3\underline{6}9 = 4.37 \text{ L}$$

5.55   Solve the ideal gas law for temperature in K, and convert to °C:

$$T = \frac{PV}{nR} = \frac{(3.50 \text{ atm})(4.00 \text{ L})}{(0.410 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})} = 41\underline{6}.1 = 416 \text{ K}$$

$$°C = 416 - 273 = 143°C$$

5.56   Find the moles of C_3H_8, solve the ideal gas law for temperature in K, and convert to °C:

$$n = 5.65 \text{ g} \times \frac{1 \text{ mol C}_3\text{H}_8}{44.09 \text{ g C}_3\text{H}_8} = 0.12\underline{8}14 \text{ mol}$$

$$T = \frac{PV}{nR} = \frac{(741/760 \text{ atm})(2.50 \text{ L})}{(0.12814 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})} = 23\underline{1}.79 \text{ K}$$

$$°C = 23\underline{1}.79 - 273 = -4\underline{1}.20 = -41°C$$

5.57   Because density equals mass per unit volume, calculating the mass of 1 L (exact
        number) of gas will give the density of the gas. Start from the ideal gas law, and
        calculate n; then convert the moles of gas to grams using the molar mass.

$$n = \frac{PV}{RT} = \frac{(751/760 \text{ atm})(1 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(304 \text{ K})} = 0.03961 \text{ mol}$$

$$0.03961 \text{ mol} \times \frac{17.03 \text{ g}}{1 \text{mol}} = 0.67458 = 0.675 \text{ g}$$

Therefore, the density of NH_3 at 31°C is 0.675 g/L.

5.58   As in the previous problem, calculate the mass of 1 L (exact number) of gas to obtain
        the density of the gas. Start from the ideal gas law, and calculate n; then convert to
        grams.

$$n = \frac{PV}{RT} = \frac{(967/760 \text{ atm})(1 \text{L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(329 \text{ K})} = 0.04712 \text{ mol}$$

$$0.04712 \text{ mol} \times \frac{34.08 \text{ g}}{1 \text{ mol}} = 1.6058 \text{ g}$$

Therefore, the density of H_2S at 56°C is 1.61 g/L.

5.59   Calculate the mass of 1 L (exact number) using the ideal gas law; convert to mass.

$$n = \frac{PV}{RT} = \frac{(0.897 \text{ atm})(1 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(297 \text{ K})} = 0.036804 \text{ mol}$$

$$0.036804 \text{ mol} \times \frac{58.12 \text{ g}}{1 \text{mol}} = 2.139 = 2.14 \text{ g}$$

Therefore, the density of C_4H_{10} is 2.14 g/L.

5.60   Calculate the mass of 1 L (exact number) using the ideal gas law; convert to mass.

$$n = \frac{PV}{RT} = \frac{(797/760 \text{ atm})(1 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(371 \text{ K})} = 0.034446 \text{ mol}$$

$$0.034446 \text{ mol} \times \frac{119.5 \text{ g}}{1 \text{ mol}} = 4.116 = 4.12 \text{ g}$$

Therefore, the density of $CHCl_3$ at 99°C is 4.12 g/L.

5.61 The ideal gas law gives n moles, which then are divided into the mass of 1.585 g for molar mass.

$$n = \frac{PV}{RT} = \frac{(753/760 \text{ atm})(1 \text{L})}{(0.08206 \text{ L} \cdot \text{atm}/ \text{K} \cdot \text{mol})(363 \text{ K})} = 0.033\underline{2}62 \text{ mol}$$

$$\text{Molar mass} = \frac{1.585 \text{ g}}{0.033262 \text{ mol}} = 47.\underline{6}51 \text{ g/mol}$$

The molecular weight is 47.7 amu.

5.62 The moles in 237 mL (0.237 L) of the compound are obtained from the ideal gas law. Dividing the mass (0.548 g) of the gas by the moles gives the molar mass and molecular weight.

$$n = \frac{PV}{RT} = \frac{(755/760 \text{ atm})(0.237 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm}/ \text{K} \cdot \text{mol})(373 \text{ K})} = 0.0076\underline{9}2 \text{ mol}$$

$$\text{Molar mass} = \frac{0.548 \text{ g}}{0.007692 \text{ mol}} = 71.\underline{2}4 \text{ g/mol}$$

The molecular weight is 71.2 amu.

5.63 The moles in 250 mL (0.250 L) of the compound are obtained from the ideal gas law. The 2.56 g mass of the gas then is divided by the moles to obtain the molar mass and molecular weight.

$$n = \frac{PV}{RT} = \frac{(786/760 \text{ atm})(0.250 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm}/ \text{K} \cdot \text{mol})(394 \text{ K})} = 0.0079\underline{9}6 \text{ mol}$$

$$\text{Molar mass} = \frac{2.56 \text{ g}}{0.007996 \text{ mol}} = 3.2\underline{0}1 \times 10^2 = 3.20 \times 10^2 \text{ g/mol}$$

The molecular weight is 3.20×10^2 amu.

5.64   The moles in 345 mL (0.345 L) of the compound are obtained from the ideal gas law. The 2.30 g mass of the gas then is divided by the moles to obtain the molar mass and molecular weight.

$$n = \frac{PV}{RT} = \frac{(985/760 \text{ atm})(0.345 \text{ L})}{(0.08206 \text{ L} \bullet \text{atm/ K} \bullet \text{mol})(421\text{K})} = 0.0129\underline{4} \text{ mol}$$

$$\text{Molar mass} = \frac{2.30 \text{ g}}{0.01294 \text{ mol}} = 17\underline{7}.7 \text{ g/mol}$$

The molecular weight is 178 amu.

5.65   For a gas at a given temperature and pressure, the density depends on molecular weight (or for a mixture, the average molecular weight). Thus, at the same temperature and pressure, the density of NH_4Cl gas would be greater than that of a mixture of NH_3 and HCl because the average molecular weight of NH_3 and HCl would be lower than that of NH_4Cl.

5.66   For a gas at a given temperature and pressure, the density depends on molecular weight (or for a mixture, the average molecular weight). Thus, at the same temperature and pressure, the density of PCl_5 gas would be greater than that of a mixture of PCl_3 and Cl_2 formed by the decomposition of PCl_5 because the average molecular weight of PCl_3 and Cl_2 would be lower than that of PCl_5.

5.67   The 0.050 mol CaC_2 will form 0.025 mol C_2H_2. The volume is found from the ideal gas law:

$$\text{Vol} = \frac{(0.050 \text{ mol})(0.08206 \text{ L} \bullet \text{atm/K} \bullet \text{mol})(299 \text{ K})}{684/760 \text{ atm}} = 1.\underline{3}6 = 1.4 \text{ L}$$

5.68   The 0.0840 mol Mg will form 0.0840 mol of H_2. The volume is found from the ideal gas law:

$$\text{Vol} = \frac{(0.0840 \text{ mol})(0.08206 \text{ L} \bullet \text{atm/K} \bullet \text{mol})(306 \text{ K})}{665/760 \text{ atm}} = 2.4\underline{1}0 = 2.41 \text{ L}$$
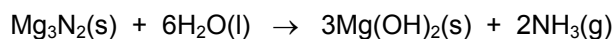
5.69   Use the equation to obtain the moles of CO_2 and then the ideal gas law to obtain the volume.

$$2LiOH(s) + CO_2(g) \rightarrow Li_2CO_3(aq) + H_2O(l)$$
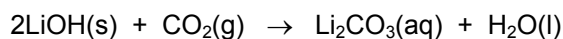
$$327 \text{ g LiOH} \times \frac{1 \text{ mol LiOH}}{23.95 \text{ g LiOH}} \times \frac{1 \text{ mol CO}_2}{2 \text{ mol LiOH}} = 6.8\underline{2}67 \text{ mol CO}_2$$

$$V = \frac{(6.8\underline{2}67 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(294 \text{ K})}{781/760 \text{ atm}} = 16\underline{0}.2 = 160. \text{ L}$$

5.70   Use the equation to obtain the moles of ammonia and then the ideal gas law to obtain the volume.

$$Mg_3N_2(s) + 6H_2O(l) \rightarrow 3Mg(OH)_2(s) + 2NH_3(g)$$

$$4.56 \text{ g Mg}_3\text{N}_2 \times \frac{1 \text{ mol Mg}_3\text{N}_2}{100.9 \text{ g Mg}_3\text{N}_2} \times \frac{2 \text{ mol NH}_3}{1 \text{ mol Mg}_3\text{N}_2} = 0.0903\underline{8}6 \text{ mol NH}_3$$

$$V = \frac{(0.090386 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(297 \text{ K})}{(753/760) \text{ atm}} = 2.2\underline{2}3 = 2.22 \text{ L}$$
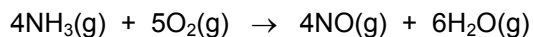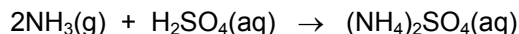
5.71   Use the equation to obtain the moles of ammonia and then the ideal gas law to obtain the volume.

$$2NH_3(g) + CO_2(g) \rightarrow NH_2CONH_2(aq) + H_2O(l)$$

$$908 \text{ g urea} \times \frac{1 \text{ mol urea}}{60.06 \text{ g urea}} \times \frac{2 \text{ mol NH}_3}{1 \text{ mol urea}} = 30.\underline{2}3 \text{ mol NH}_3$$

$$V = \frac{(30.\underline{2}3 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(298 \text{ K})}{3.00 \text{ atm}} = 24\underline{6}.4 = 246 \text{ L}$$

5.72 Use the equation to obtain the moles of ammonia and then the ideal gas law to obtain the volume.

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$$

$$50.0 \text{ g NO} \times \frac{1 \text{ mol NO}}{30.0 \text{ g NO}} \times \frac{5 \text{ mol O}_2}{4 \text{ mol NO}} = 2.0\underline{8}3 \text{ mol O}_2$$

$$V = \frac{(2.083 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(308 \text{ K})}{2.15 \text{ atm}} = 24.\underline{4}87 = 24.5 \text{ L}$$

5.73 Use the equation to obtain the moles of ammonia and then the ideal gas law to obtain the volume.

$$2NH_3(g) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq)$$

$$150.0 \text{ g } (NH_4)_2SO_4 \times \frac{1 \text{ mol } (NH_4)_2SO_4}{132.1 \text{ g } (NH_4)_2SO_4} \times \frac{2 \text{ mol NH}_3}{1 \text{ mol } (NH_4)_2SO_4}$$

$$= 2.2\underline{7}10 \text{ mol NH}_3$$

$$V = \frac{nRT}{P} = \frac{(2.2710 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(288 \text{ K})}{1.15 \text{ atm}} = 46.\underline{6}7 = 46.7 \text{ L}$$

5.74 Use the equation to obtain the moles of ammonia, and then the ideal gas law to obtain the volume.

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$$
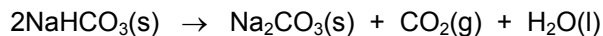
$$26.8 \text{ g NaHCO}_3 \times \frac{1 \text{ mol NaHCO}_3}{84.0 \text{ g NaHCO}_3} \times \frac{1 \text{ mol CO}_2}{2 \text{ mol NaHCO}_3} = 0.159\underline{5}2 \text{ mol CO}_2$$

$$V = \frac{nRT}{P} = \frac{(0.15952 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(350 \text{ K})}{(756/760) \text{ atm}} = 4.6\underline{0}59 = 4.61 \text{ L}$$

5.75   Calculate the partial pressure of each gas; then add the pressures since the total pressure is equal to the sum of the partial pressures:

$$P(He) = \frac{(0.0200 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(283 \text{ K})}{2.50 \text{ L}} = 0.18\underline{5}78 \text{ atm}$$

$$P(He) = \frac{(0.0100 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(283 \text{ K})}{2.50 \text{ L}} = 0.092\underline{8}9 \text{ atm}$$

The total pressure = 0.18587 = 0.092\underline{8}9 = 0.27\underline{8}67 = 0.279

5.76   Calculate the partial pressure of each gas; then add the pressures since the total pressure is equal to the sum of the partial pressures:

$$P(He) = \frac{(0.0300 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(293 \text{ K})}{4.00 \text{ L}} = 0.18\underline{0}3 \text{ at m}$$

$$P(O_2) = \frac{(0.0200 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(293 \text{ K})}{4.00 \text{ L}} = 0.12\underline{0}2 \text{ atm}$$

The total pressure = 0.18\underline{0}3 + 0.12\underline{0}2 = 0.30\underline{0}5 = 0.301 atm

5.77   Convert mass of O_2 and mass of He to moles. Use the ideal gas law to calculate the partial pressures, and then add to obtain the total pressures.

$$0.00103 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} = 3.2\underline{1}9 \times 10^{-5} \text{ mol } O_2$$

$$0.00056 \text{ g He} \times \frac{1 \text{ mol He}}{4.00 \text{ g He}} = 1.\underline{4}0 \times 10^{-4} \text{ mol}$$

$$P = \frac{nRT}{V} = \frac{(3.219 \times 10^{-5} \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(288 \text{ K})}{0.2000 \text{ L}}$$

$$= 0.00380\underline{4} \text{ atm } O_2$$

$$P = \frac{nRT}{V} = \frac{(1.\underline{4}0 \times 10^{-4} \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(288 \text{ K})}{0.2000 \text{ L}} = 0.01\underline{6}54 \text{ atm He}$$

$$P = P_{O_2} + P_{He} = 0.00380\underline{4} \text{ atm} + 0.01\underline{6}54 \text{ atm} = 0.02\underline{0}3 = 0.020 \text{ atm}$$

5.78   Obtain the partial pressure of helium using Dalton's law; then use the ideal gas law to obtain the mass of helium.

$$P_{He} = P - P_{O_2} = 3.00 \text{ atm} - 0.200 \text{ atm} = 2.8\underline{0}0 \text{ atm}$$

$$n_{He} = \frac{P_{He}V}{RT} = \frac{(2.80 \text{ atm})(10.0 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(293 \text{ K})} = 1.1\underline{6}45 \text{ mol He}$$

$$Mass_{He} = 1.1645 \text{ mol He} \times \frac{4.003 \text{ g He}}{1 \text{ mol He}} = 4.6\underline{6}1 = 4.66 \text{ g He}$$

5.79   For each gas, P(gas) = P x (mole fraction of gas).

$$P(H_2) = 760 \text{ mmHg} \times 0.250 = 19\underline{0}.0 = 190 \text{ mmHg}$$

$$P(CO_2) = 760 \text{ mmHg} \times 0.650 = 49\underline{4}.0 = 494 \text{ mmHg}$$

$$P(HCl) = 760 \text{ mmHg} \times 0.054 = 4\underline{1}.04 = 41 \text{ mmHg}$$

$$P(HF) = 760 \text{ mmHg} \times 0.028 = 2\underline{1}.28 = 21 \text{ mmHg}$$

$$P(SO_2) = 760 \text{ mmHg} \times 0.017 = 1\underline{2}.92 = 13 \text{ mmHg}$$

$$P(H_2S) = 760 \text{ mmHg} \times 0.001 = 0.\underline{7}6 = 0.8 \text{ mmHg}$$

5.80   For each gas, P(gas) = P x (mole fraction of gas).

$$P(He) = 6.91 \text{ atm} \times 0.790 = 5.4\underline{5}89 = 5.46 \text{ atm}$$
$$P(N_2) = 6.91 \text{ atm} \times 0.170 = 1.1\underline{7}47 = 1.17 \text{ atm}$$
$$P(O_2) = 6.91 \text{ atm} \times 0.040 = 0.2\underline{7}64 = 0.28 \text{ atm}$$

5.81   The total pressure is the sum of the partial pressures of CO and H_2O, so

$$P_{CO} = P - P_{water} = 689 \text{ mmHg} - 23.8 \text{ mmHg} = 66\underline{5}.2 \text{ mmHg}$$

$$n_{CO} = \frac{P_{CO}V}{RT} = \frac{(665/760 \text{ atm})(3.85 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/ K} \cdot \text{mol})(298 \text{ K})} = 0.13\underline{7}8 \text{ mol CO}$$

$$0.13\underline{7}8 \text{ g CO} \times \frac{1 \text{ mol HCOOH}}{1 \text{ mol CO}} \times \frac{46.03 \text{ g HCOOH}}{1 \text{ mol HCOOH}}$$

$$= 6.3\underline{4}2 = 6.34 \text{ g HCOOH}$$

5.82 The total pressure is the sum of the partial pressures of N_2 and H_2O, so

$$P_{N_2} = P - P_{H_2O} = 97.8 \text{ kPa} \times \left(\frac{1 \text{ atm}}{101.3 \text{ kPa}} \right) - 16.5 \text{ mmHg} \times \left(\frac{1 \text{ atm}}{760 \text{ mmHg}} \right)$$

$$= 0.94\underline{3}7 \text{ atm}$$

$$n_{N_2} = \frac{P_{N_2} V}{RT} = \frac{(0.94\underline{3}7 \text{ atm}) (3.75 \text{ L})}{0.08206 \text{ L} \cdot \text{atm} / \text{K} \cdot \text{mol}) (299 \text{ K})} = 0.14\underline{4}2 \text{ mol } N_2$$

$$0.14\underline{4}2 \text{ mol } N_2 \times \frac{1 \text{ mol } NH_4NO_2}{1 \text{ mol } N_2} \times \frac{64.04 \text{ g } NH_4NO_2}{1 \text{ mol } NH_4NO_2} = 9.2\underline{3}7 = 9.24 \text{ g } NH_4NO_3$$

5.83 Substitute 298 K (25°C) and 398 K (125°C) into Maxwell's distribution:

$$u_{25} = \sqrt{\frac{3RT}{M_m}} = \sqrt{\frac{3 \times 8.31 \text{ kg} \cdot \text{m}^2/(\text{s}^2 \cdot \text{K} \cdot \text{mol}) \times 298 \text{ K}}{28.02 \times 10^{-3} \text{ kg/mol}}}$$

$$= 5.1\underline{4}9 \times 10^2 = 5.15 \times 10^2 \text{ m/s}$$

$$u_{125} = \sqrt{\frac{3RT}{M_m}} = \sqrt{\frac{3 \times 8.31 \text{ kg} \cdot \text{m}^2/(\text{s}^2 \cdot \text{K} \cdot \text{mol}) \times 398 \text{ K}}{28.02 \times 10^{-3} \text{ kg/mol}}}$$

$$= 5.9\underline{5}07 \times 10^2 = 5.95 \times 10^2 \text{ m/s}$$

Graph as in Figure 5.23.

5.84 Substitute 296 K (23°C) into Maxwell's distribution:

$$u_{23} = \sqrt{\frac{3RT}{M_m}} = \sqrt{\frac{3 \times 8.31 \text{ kg} \cdot \text{m}^2/(\text{s}^2 \cdot \text{K} \cdot \text{mol}) \times 296 \text{ K}}{159.8 \times 10^{-3} \text{ kg/mol}}} = 21\underline{4}.8 = 215 \text{ m/s}$$

The rms speed is the same at 1.50 atm as at 1.00 atm because the temperature is the same.

5.85 Substitute 330 K (57°C) into Maxwell's distribution:

$$u_{330\ K} = \sqrt{\frac{3RT}{M_m}} = \sqrt{\frac{3 \times 8.31\ kg \cdot m^2/(s^2 \cdot K \cdot mol) \times 330\ K}{352 \times 10^{-3}\ kg/mol}}$$

$$= 1.5\underline{2}8 \times 10^2 = 1.53 \times 10^2\ m/s$$

5.86 Substitute 365 K into Maxwell's distribution:

$$u_{365\ K} = \sqrt{\frac{3RT}{M_m}} = \sqrt{\frac{3 \times 8.31\ kg \cdot m^2/(s^2 \cdot K \cdot mol) \times 365\ K}{2.016 \times 10^{-3} kg/mol}}$$

$$= 2.1\underline{2}4 \times 10^3 = 2.12 \times 10^3\ m/s$$

This is equal to 2.12 km/s, which is less than the escape velocity.

5.87 Because $u(CO_2) = u(H_2)$, we can equate the two right-hand sides of the Maxwell distributions:

$$\sqrt{\frac{3RT(CO_2)}{M_m(CO_2)}} = \sqrt{\frac{3RT(H_2)}{M_m(H_2)}}$$

Squaring both sides, rearranging to solve for $T(CO_2)$, and substituting numerical values, we have

$$T(CO_2) = T(H_2) \times \frac{M_m(CO_2)}{M_m(H_2)} = 298\ K \times \frac{44.01\ g/mol}{2.016\ g/mol} = 65\underline{0}5.4\ K$$

Thus the temperature is

$$T = 65\underline{0}5.4 - 273 = 62\underline{3}2 = 6.23 \times 10^3\ °C$$

5.88 Substitute 375 m/s into the Maxwell distribution, square both sides, and solve for T:

$$u = 400\ m/s = \sqrt{\frac{3 \times 8.31\ kg \cdot m^2/(s^2 \cdot K \cdot mol) \times T}{32.00 \times 10^{-3}\ kg/mol}}$$

$$(400)^2 = 160000 = \frac{3 \times 8.31 \times T}{32.00 \times 10^{-3}\ K}$$

$$T = 20\underline{5}.37\ K = -67.\underline{7}7 = -67.8\ °C$$

5.89 Because the ratio is the same at any temperature, $T(N_2) = T(O_2)$. Write a ratio of two Maxwell distributions after omitting $T(O_2)$ and $T(H_2)$ in each distribution:

$$\frac{u_{N_2}}{u_{O_2}} = \frac{\sqrt{\dfrac{3RT(N_2)}{M_m}}}{\sqrt{\dfrac{3RT(O_2)}{M_m}}} = \frac{\sqrt{\dfrac{3 \times 8.31 \text{ kg} \cdot \text{m}^2/(\text{s}^2 \cdot \text{K} \cdot \text{mol})}{28.02 \times 10^{-3} \text{ kg/mol}}}}{\sqrt{\dfrac{3 \times 8.31 \text{ kg} \cdot \text{m}^2/(\text{s}^2 \cdot \text{K} \cdot \text{mol})}{32.00 \times 10^{-3} \text{ kg/mol}}}}$$

$$\frac{u_{N_2}}{u_{O_2}} = \sqrt{\frac{32.00}{28.02}} = \frac{1.06\underline{8}6}{1} = \frac{1.069}{1}$$

5.90 Because the ratio is the same at any temperature, $T(H_2) = T(H_2Te)$. Write a ratio of two Maxwell distributions after omitting $T(H_2)$ and $T(H_2Te)$ in each distribution:

$$\frac{u_{H_2}}{u_{H_2Te}} = \frac{\sqrt{\dfrac{3RT(H_2)}{M_m}}}{\sqrt{\dfrac{3RT(H_2Te)}{M_m}}} = \frac{\sqrt{\dfrac{3 \times 8.31 \text{ kg} \cdot \text{m}^2/(\text{s}^2 \cdot \text{K} \cdot \text{mol})}{2.016 \times 10^{-3} \text{ kg/mol}}}}{\sqrt{\dfrac{3 \times 8.31 \text{ kg} \cdot \text{m}^2/(\text{s}^2 \cdot \text{K} \cdot \text{mol})}{129.62 \times 10^{-3} \text{ kg/mol}}}}$$

$$\frac{u_{H_2}}{u_{H_2Te}} = \sqrt{\frac{129.62}{2.016}} = \frac{8.0\underline{1}8}{1} = \frac{8.02}{1}$$

5.91 Because the ratio is the same at any temperature, $T(H_2) = T(I_2)$. A ratio of two Maxwell distributions can be written as in the previous two problems, but this also can be simplified by canceling the 3×8.31 terms and rearranging the denominators to give

$$\frac{u_{H_2}}{u_{I_2}} = \frac{\sqrt{M_m(I_2)}}{\sqrt{M_m(H_2)}} = \sqrt{\frac{253.8}{2.016}} = \frac{11.2\underline{2}05}{1} = \frac{11.22}{1}$$

Because hydrogen diffuses 11.22 times as fast as iodine, the time it would take would be 1/11.22 of the time required for iodine:

$$t(H_2) = 39 \text{ s} \times (1/11.22) = 3.\underline{4}7 = 3.5 \text{ s}$$

5.92 Because the ratio is the same at any temperature, T(He) = T(N$_2$). A ratio of two Maxwell distributions can be written, but this also can be simplified by canceling the 3 x 8.31 terms and rearranging the denominators to give

$$\frac{u_{He}}{u_{N_2}} = \frac{\sqrt{M_m(N_2)}}{\sqrt{M_m(He)}} = \sqrt{\frac{28.02}{4.00}} = \frac{2.6\underline{4}7}{1} = \frac{2.65}{1}$$

Because helium diffuses 2.65 times as fast as nitrogen, the time it would take would be 1/2.65 of the time required for nitrogen:

t(He) = 10.6 hr x (1/2.65) = 4.00 hr

5.93 Because the diffusion occurs at the same temperature, T(gas) = T(Ar). A ratio of two Maxwell distributions can be written, but it can be simplified by canceling the 3 x 8.31 terms and rearranging the denominators. To simplify the definition of the rates, we assume the time is one second, and define the rate(Ar) as 9.23 mL/1 s and the rate(gas) as 4.83 mL/1 s. Then we write a ratio of two Maxwell distributions:

$$\frac{u_{gas}}{u_{Ar}} = \frac{4.83 \text{ mL}/1 \text{ s}}{9.23 \text{ mL}/1 \text{ s}} = \frac{\sqrt{M_m(Ar)}}{\sqrt{M_m(gas)}} = \frac{\sqrt{39.95}}{\sqrt{M_m(gas)}}$$

Squaring both sides gives

$$0.27\underline{3}83 = \frac{39.95}{M_m(gas)}$$

Solving for the molar mass gives

$$M_m(gas) = \frac{39.95}{0.27383} = 14\underline{5}.8 \text{ g/mol}; \text{ molecular weight} = 146 \text{ amu}$$

5.94 As in the previous problem, the temperatures are the same and a ratio of two Maxwell distributions can be written and simplified. In contrast, here we assume the volume is one mL, and we define the rate(N$_2$) as 1 mL/68.3 s and the rate(gas) as 1 mL/85.6 s. Then we write

$$\frac{u_{gas}}{u_{N_2}} = \frac{1 \text{ mL}/85.6 \text{ s}}{1 \text{ mL}/68.3 \text{ s}} = \frac{\sqrt{M_m(N_2)}}{\sqrt{M_m(gas)}} = \frac{\sqrt{28.02}}{\sqrt{M_m(gas)}}$$

Squaring both sides gives

$$0.63\underline{6}639 = \frac{28.02}{M_m(gas)}$$

Solving for the molar mass gives

$$M_m(gas) = \frac{28.02}{0.636639} = 44.\underline{0}1 \text{ g/mol; molecular weight} = 44.0 \text{ amu}$$

5.95   Solving the van der Waals equation for n = 1 and T = 355.2 K for P gives

$$P = \frac{RT}{(V - b)} - \frac{a}{V^2} = \frac{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(355.2 \text{ K})}{(30.00 \text{ L} - 0.08710 \text{ L})} - \frac{12.56 \text{ L}^2 \cdot \text{atm}}{(30.00 \text{ L})^2}$$

P = 0.974$\underline{4}$19 - 0.0139$\underline{5}$55 = 0.960$\underline{4}$6 = 0.9605 atm

P(ideal gas law) = 0.971$\underline{5}$9 atm

5.96   Solving the van der Waals equation for n = 1 and T = 393.2 K for P gives

$$P = \frac{RT}{(V - b)} - \frac{a}{V^2} = \frac{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(393.2 \text{ K})}{(32.50 \text{ L} - 0.03049 \text{ L})} - \frac{5.537 \text{ L}^2 \cdot \text{atm}}{(32.50 \text{ L})^2}$$

P = 0.993$\underline{7}$3 - 0.005242$\underline{1}$ = 0.988$\underline{4}$90 = 0.9885 atm

P(ideal gas law) = 0.992$\underline{7}$9 atm

5.97   To calculate a/V^2 in the van der Waals equation, we obtain V from the ideal gas law at 1.00 atm:

$$V = \frac{RT}{P} = \frac{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(273 \text{ K})}{1.00 \text{ atm}} = 22.40 \text{ L}$$

$$\frac{a}{V^2} = \frac{5.570}{(22.4 \text{ L})^2} = 1.1\underline{1}0 \times 10^{-2}$$

At 1.00 atm, V = 22.4 L, and a/V^2 = 1.110 x 10^{-2}. Substituting into the van der Waals equation:

$$V = \frac{RT}{P + \dfrac{a}{V^2}} + b = \frac{(0.08206 \text{ L} \bullet \text{atm/K} \bullet \text{mol})(273 \text{ K})}{(1.00 + 1.110 \text{ x } 10^{-2})\text{atm}} + 0.06499$$

$$= 22.\underline{2}2 = 22.2 \text{ L}$$

At 10.0 atm, the van der Waals equation gives 2.08 L. The ideal gas law gives 22.4 L for 1.00 atm and 2.24 L for 10.0 atm.

5.98　To substitute for a/V^2 in the van der Waals equation, we obtain V from the ideal gas law at 1.00 atm:

$$V = \frac{RT}{P} = \frac{(0.08206 \text{ L} \bullet \text{atm/K} \bullet \text{mol})(273 \text{ K})}{1.00 \text{ atm}} = 22.40 \text{ L}$$

$$\frac{a}{V^2} = \frac{1.382}{(22.4 \text{ L})^2} = 2.7\underline{5}4 \text{ x } 10^{-3}$$

At 1.00 atm, V = 22.4 L, and a/V^2 = 2.75 x 10^{-3}. Substituting into the van der Waals equation:

$$V = \frac{RT}{P + \dfrac{a}{V^2}} + b = \frac{(0.08206 \text{ L} \bullet \text{atm/K} \bullet \text{mol})(273 \text{ K})}{(1.00 + 2.75 \text{ x } 10^{-3})\text{atm}} + 0.03186$$

$$= 22.\underline{3}7 = 22.4 \text{ L}$$

At 10.0 atm, the van der Waals equation gives 2.21 L. The ideal gas law gives 22.4 L for 1.00 atm and 2.24 L for 10.0 atm.

■ Solutions to General Problems

5.99 Calculate the mass of one cm^2 of the 20.5 m of water above the air in the glass. The volume is the product of the area of one cm^2 and the height of 20.5×10^2 cm (20.5 m) of water. The density of 1.00 g/cm^3 must be used to convert volume to mass:

$$m = d \times V$$

$$m = 1.00 \text{ g/cm}^3 \times (1.00 \text{ cm}^2 \times 20.5 \times 10^2 \text{ cm}) = 2.05 \times 10^3 \text{ g, or } 2.05 \text{ kg}$$

The pressure exerted on an object at the bottom of the column of water is

$$P = \frac{\text{force}}{\text{area}} = \frac{(m)(g)}{\text{area}} = \frac{(2.05 \text{ kg})(9.807 \text{ m/s}^2)}{(1.00 \text{ cm}^2)\left(\frac{10^{-2} \text{ m}}{1 \text{cm}}\right)^2} = 2.01 \times 10^5 \text{ kg/ms}^2$$

$$= 2.01 \times 10^5 \text{ Pa}$$

The total pressure on the air in the tumbler equals the barometric pressure and the water pressure:

$$P = 1.00 \times 10^2 \text{ kPa} + 2.01 \times 10^2 \text{ kPa} = 3.01 \times 10^2 \text{ kPa}$$

Multiply the initial volume by a factor accounting for the change in pressure to find V_f:

$$V_f = V_i \times \frac{P_i}{P_f} = 243 \text{ cm}^3 \times \left(\frac{1.00 \times 10^2 \text{ kPa}}{3.01 \times 10^2 \text{ kPa}}\right) = 80.\underline{7}3 = 80.7 \text{ cm}^3$$

5.100 The volume of one m^2 of the 30.0 m of water above the air is 30.0 m^3. The mass of water above this area is:

$$m = d \times V = \frac{1.025 \text{ g}}{1 \text{cm}^3} \times \left(\frac{1 \text{cm}}{10^{-2} \text{ m}}\right)^3 \times 30.0 \text{ m}^3$$

$$= 3.0\underline{7}5 \times 10^7 \text{ g} \ (3.0\underline{7}5 \times 10^4 \text{ kg})$$

The pressure on this area is

$$P = \frac{\text{force}}{\text{area}} = \frac{(m)(g)}{\text{area}} = \frac{(3.075 \times 10^4 \text{ kg})(9.807 \text{ m/s}^2)}{1 \text{m}^2}$$

$$= 3.0\underline{1}56 \times 10^5 \text{ kg/ms}^2 \text{ or Pa}$$

$$P = 3.0156 \times 10^5 \text{ Pa} \times \frac{1 \text{ atm}}{1.013 \times 10^5 \text{ Pa}} = 2.9\underline{7}69 \text{ atm}$$

The total pressure at 30.0 m is

$$P = P_{air} + P_{water} = 1.00 \text{ atm} + 2.9769 \text{ atm} = 3.9\underline{7}69 \text{ atm}$$

The density of a gas is directly proportional to the pressure, so

$$d_2 = d_1 \times \frac{P_2}{P_1} = (1.205 \text{ g/L}) \times \left(\frac{3.9769 \text{ atm}}{1.00 \text{ atm}}\right) = 4.7\underline{9}2 = 4.79 \text{ g/L}$$

5.101 Use the combined gas law and solve for V_f :

$$V_f = V_i \times \frac{P_i}{P_f} \times \frac{T_f}{T_i} = 201 \text{ mL} \times \frac{738 \text{ mmHg}}{760 \text{ mmHg}} \times \frac{273 \text{ K}}{294 \text{ K}}$$

$$= 18\underline{1}.24 = 181 \text{ mL}$$

5.102 Use the combined gas law and solve for V_f:

$$V_f = V_i \times \frac{P_i}{P_f} \times \frac{T_f}{T_i} = 12.0 \text{ L} \times \frac{11.0 \text{ atm}}{1.00 \text{ atm}} \times \frac{273 \text{ K}}{293 \text{ K}} = 12\underline{2}.98 = 123 \text{ L}$$

5.103 Use the combined gas law and solve for V_f :

$$V_f = V_i \times \frac{P_i}{P_f} \times \frac{T_f}{T_i} = 5.0 \text{ dm}^3 \times \frac{100.0 \text{ kPa}}{79.0 \text{ kPa}} \times \frac{293 \text{ K}}{287 \text{ K}} = 6.\underline{4}6 = 6.5 \text{ dm}^3$$

5.104 Let $V_i = 1$ volume (vol) in the combined gas law; solving for V_f will give the relative final volume, or factor, for the increase:

$$V_f = V_i \times \frac{P_i}{P_f} \times \frac{T_f}{T_i} = 1 \text{ vol} \times \frac{1.00 \text{ atm}}{1.00 \times 10^{-3} \text{ atm}} \times \frac{252 \text{ K}}{288 \text{ K}}$$

$$= 87\underline{5}.0 \text{ vol, or 875 times}$$

5.105   Use the ideal gas law to calculate the moles of helium, and combine this with Avogadro's number to obtain the number of helium atoms:

$$n = \frac{PV}{RT} = \frac{(765/760 \text{ atm})(0.01205 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm}/ \text{K} \cdot \text{mol})(296 \text{ K})} = 4.9\underline{9}3 \times 10^{-4} \text{ mol}$$

$$4.993 \times 10^{-4} \text{ mol He} \times \frac{6.022 \times 10^{23} \text{ He}^{2+} \text{ ions}}{1 \text{ mol He}} \times \frac{1 \text{ atom}}{1 \text{ He}^{2+} \text{ ion}} = 3.0\underline{0}67 \times 10^{20}$$

$$= 3.01 \times 10^{20} \text{ atoms}$$

5.106   Use the ideal gas law to calculate the moles of nitrogen, and combine this with the molar mass to obtain the mass of nitrogen. Then calculate the mass percentage:

$$n = \frac{PV}{RT} = \frac{(749/760 \text{ atm})(0.00159 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm}/\text{K} \cdot \text{mol})(298 \text{ K})} = 6.4\underline{0}7 \times 10^{-5} \text{ mol}$$

$$6.4\underline{0}7 \times 10^{-5} \text{ mol N}_2 \times \frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} = 1.7\underline{9}55 \times 10^{-3} \text{ g N} \ (1.7\underline{9}55 \text{ mg N})$$

$$\text{Percent N} = \frac{\text{mass N}}{\text{mass comp.}} \times 100 = \frac{1.7955 \text{ mg N}}{8.75 \text{ mg}} \times 100\% = 20.\underline{5}2 = 20.5\%$$

5.107   Calculate the molar mass, M_m, by dividing the mass of one L of air by the moles of the gas from the ideal gas equation:

$$M_m = \frac{\text{mass}}{n} = 1.2929 \text{ g air} \times \frac{(0.082058 \text{ L} \cdot \text{atm}/\text{K} \cdot \text{mol})(273.15 \text{ K})}{(1 \text{atm})(1 \text{L})}$$

$$= 28.97\underline{9}2 \text{ g/mol} = 28.979 \text{ g/mol (amu)}$$

5.108 First, calculate the molar mass, M_m, by dividing the mass of one L of air by the moles of the gas from the ideal gas equation:

$$M_m = \frac{mass}{n} = 1.22 \text{ g gas} \times \frac{(0.08206 \text{ L} \bullet \text{atm}/\text{K} \bullet \text{mol})(293 \text{ K})}{(1 \text{atm})(1 \text{L})} = 29.\underline{3}3 \text{ g/mol}$$

Find the empirical formula from the 80.0 percent C and 20 percent H by assuming one gram of compound and calculating the moles:

$$0.800 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.06661 \text{ mol C}$$

$$0.200 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.1984 \text{ mol H}$$

$$\frac{0.06661 \text{ mol C}}{0.06661} = 1.00, \text{ or 1 mol C;} \quad \frac{0.1984 \text{ mol H}}{0.06661} = 2.98, \text{ or 3 mol H}$$

The simplest formula is CH_3, whose empirical formula weight is 15.03 g/mol. The number of CH_3 units contained in the molecular weight of 29.33 is

$$\frac{\text{molecular weight}}{\text{formula weight}} = \frac{29.33}{15.03} = 1.95, \text{ or 2}$$

The molecular formula is $(CH_3)_2$, or C_2H_6.

5.109 Use the ideal gas law to calculate the moles of CO_2. Then convert to mass of LiOH.

$$n = \frac{PV}{RT} = \frac{(1.00 \text{ atm})(5.8 \times 10^2 \text{ L})}{(0.08206 \text{ L} \bullet \text{atm}/\text{K} \bullet \text{mol})(273 \text{ K})} = 2\underline{5}.89 \text{ mol CO}_2$$

$$2\underline{5}.89 \text{ mol CO}_2 \times \frac{2 \text{ mol LiOH}}{1 \text{ mol CO}_2} \times \frac{23.95 \text{ g LiOH}}{1 \text{ mol LiOH}} = 1\underline{2}40 = 1.2 \times 10^3 \text{ g LiOH}$$

5.110 Use the ideal gas law to calculate moles of CO_2. Then convert to mass of pyruvic acid.

$$n = \frac{PV}{RT} = \frac{(349/760 \text{ atm})(0.0203 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(303 \text{ K})} = 3.7\underline{4}9 \times 10^{-4} \text{ mol } CO_2$$

$$3.7\underline{4}9 \times 10^{-4} \text{ mol } CO_2 \times \frac{1 \text{ mol } C_3H_4O_3}{1 \text{ mol } CO_2} \times \frac{88.06 \text{ g } C_3H_4O_3}{1 \text{ mol } C_3H_4O_3}$$

$$= 0.033\underline{0}1 = 0.0330 \text{ g } C_3H_4O_3$$

5.111 Convert mass to moles of $KClO_3$, and then use the equation below to convert to moles of O_2. Use the ideal gas law to convert moles of O_2 to pressure at 25°C (298 K).

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

$$170.0 \text{ g } KClO_3 \times \frac{1 \text{ mol } KClO_3}{122.55 \text{ g } KClO_3} \times \frac{3 \text{ mol } O_2}{2 \text{ mol } KClO_3} = 2.0\underline{8}0 \text{ mol } O_2$$

$$P = \frac{nRT}{V} = \frac{(2.0\underline{8}0 \text{ mol}) (0.08206 \text{ L} \cdot \text{atm/ K} \cdot \text{mol}) (298 \text{ K})}{2.50 \text{ L}}$$

$$= 20.\underline{3}53 = 20.4 \text{ atm } O_2$$

5.112 Convert mass to moles of $KCHO_2$, and then use the equation below to convert to moles of H_2. Use the ideal gas law to convert moles of H_2 to pressure at 21°C (294 K).

$$KCHO_2(s) + KOH(s) \rightarrow K_2CO_3(s) + H_2(g)$$
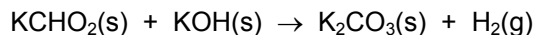
$$50.0 \text{ g } KCHO_2 \times \frac{1 \text{ mol } KCHO_2}{84.12 \text{ g } KCHO_2} = 0.59\underline{4}38 \text{ mol } KCHO_2$$

From the equation, 0.59438 mol of $KCHO_2$ produces 0.59438 mol of H_2.

$$P = \frac{nRT}{V} = \frac{(0.59438 \text{ mol } O_2)(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(294 \text{ K})}{2.50 \text{ L}}$$

$$= 5.7\underline{3}6 = 5.74 \text{ atm } H_2$$

5.113   Find the number of moles of CO_2 first. Then convert this to moles HCl and molarity of HCl.

$$n = \frac{(727 / 760 \text{ atm}) \text{ x } 0.141 \text{ L}}{0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol} \text{ x } 300 \text{ K}} = 0.0054\underline{7}88 \text{ mol}$$

mol HCl = 0.0054788 mol CO_2 x (2 mol HCl /1 mol CO_2) = 0.010\underline{9}57 mol

M HCl = 0.010\underline{9}57 mol HCl ÷ 0.0249 L HCl = 0.44\underline{00} = 0.440 mol/L HCl

5.114   Find the number of moles of CO_2 first. Then convert this to moles HCl and molarity of HCl.

$$n = \frac{731/760 \text{ atm} \text{ x } 0.159 \text{ L}}{0.08206 \text{ L} \cdot \text{atm/ K} \cdot \text{mol} \text{ x } 296 \text{ K}} = 0.00629\underline{6} \text{ mol}$$

$$\text{mol HCl} = 0.006296 \text{ mol } CO_2 \text{ x } \frac{2 \text{ mol HCl}}{1 \text{mol } CO_2} = 0.0125\underline{9} \text{ mol}$$

$$\text{Molarity} = \frac{0.01259 \text{ mol}}{0.0186 \text{ L}} = 0.677\underline{0} = 0.677 \text{ M HCl}$$

5.115   The number of moles of carbon dioxide is:

$$n = \frac{646 \ /760 \text{ atm} \text{ x } 0.1500 \text{ L}}{0.08206 \text{ L} \cdot \text{atm/ K} \cdot \text{mol} \text{ x } 300 \text{ K}} = 0.00517\underline{9} \text{ mol}$$

The number of moles of molecular acid used is:

0.1250 mol/L x 0.04141 L = 0.005176 mol acid

Thus, the acid is H_2SO_4 since one mole of H_2SO_4 reacts to form one mole of CO_2

5.116   The number of moles of carbon dioxide is:

$$n = \frac{722 \ /760 \text{ atm} \text{ x } 0.1250 \text{ L}}{0.08206 \text{ L} \cdot \text{atm/ K} \cdot \text{mol} \text{ x } 290 \text{ K}} = 0.00499\underline{0} \text{ mol}$$

The number of moles of molecular acid used is:

0.2040 mol/L x 0.04890 L = 0.0099756 mol acid

Thus, the acid is HCl since the ratio of acid to CO_2 = 0.0099756 ÷ 0.004990 = 2.00 to 1.

5.117   Use Maxwell's distribution to calculate the temperature in kelvins; then convert to °C.

$$T = \frac{u^2 M_m}{3R} = \frac{(0.605 \times 10^3 \text{ m/s})^2 \ (17.03 \times 10^{-3} \text{ kg/mol})}{3 \ (8.31 \text{ kg} \cdot \text{m}^2/\text{s}^2 \cdot \text{K} \cdot \text{mol})}$$

$$= 250.0 = 200. \text{ K} \ (-23°\text{C})$$

5.118   Use Maxwell's distribution to calculate the temperature in kelvins.

$$T = \frac{u^2 M_m}{3R} = \frac{(3.53 \times 10^3 \text{ m/s})^2 \ (4.003 \times 10^{-3} \text{ kg/mol})}{3 \ (8.31 \text{ kg} \cdot \text{m}^2/\text{s}^2 \cdot \text{K} \cdot \text{mol})}$$

$$= 2.0008 \times 10^3 = 2.00 \times 10^3 \text{ K}$$

5.119   Calculate the ratio of the root-mean-square molecular speeds, which is the same as the ratio of the rates of effusion:

$$\frac{u \text{ of } U(235)F_6}{u \text{ of } U(238)F_6} = \frac{\sqrt{M_m[U(238)F_6]}}{\sqrt{M_m[U(235)F_6]}} = \sqrt{\frac{352.04 \text{ g/mol}}{349.03 \text{ g/mol}}}$$

$$= \frac{1.004302}{1} = \frac{1.0043}{1}$$

5.120   Calculate the ratio of the root-mean-square molecular speeds, which is the same as the ratio of the rates of effusion through the barrier:

$$\frac{u \text{ of } H_2(1,1)}{u \text{ of } H_2(1,2)} = \frac{\sqrt{M_m[H_2(1,2)]}}{\sqrt{M_m[H_2(1,1)]}} = \sqrt{\frac{3.0219 \text{ g/mol}}{2.0156 \text{ g/mol}}} = \frac{1.22444}{1} = \frac{1.1224}{1}$$

5.121   First, calculate the apparent molar masses at each pressure using the ideal gas law. Only the calculation of the apparent molar mass for 0.2500 atm will be shown; the other values will be summarized in a table.

$$n = \frac{PV}{RT} = \frac{(0.2500 \text{ atm})(3.1908 \text{ L})}{(0.082057 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(273.15 \text{ K})} = 3.55896 \times 10^{-2} \text{ mol}$$
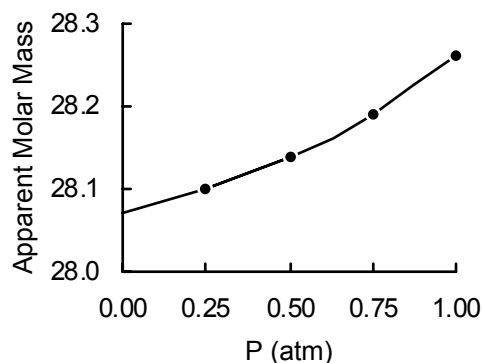
$$\text{Apparent molar mass} = \frac{1.000\,g}{3.55896 \times 10^{-2}\,mol} = 28.0\underline{9}8 = 28.10\,g/mol$$

The following table summarizes the apparent molar masses calculated as above for all P's; these data are plotted in the graph to the right of the table.

P (atm)	App. Molar Mass (g/mol)
0.2500	28.10
0.5000	28.14
0.7500	28.19
1.0000	28.26

Extrapolation back to P = 0 gives 28.07 g/mol for the molar mass of the unknown gas (CO).

5.122   First, calculate the apparent molar mass at each pressure using the ideal gas law. Only the calculation of the molar mass at 0.2500 atm will be shown; the other values will be summarized in a table.

$$n = \frac{PV}{RT} = \frac{(0.2500\,atm)(2.801\,L)}{(0.082057\,L \cdot atm/\,K \cdot mol)(273.15\,K)} = 3.12\underline{4}2 \times 10^{-2}\,mol$$

$$\text{Apparent molar mass} = \frac{1.000\,g}{3.1242 \times 10^{-2}\,mol} = 32.0\underline{0}8 = 32.01\,g/mol$$

The following table summarizes the apparent molar masses calculated as above for all P's; these data are plotted in the graph to the right of the table.

P (atm)	App. Molar Mass (g/mol)
0.2500	32.008
0.5000	32.020
0.7500	32.021
1.0000	32.029

Extrapolation back to P = 0 gives close to 32.00 g/mol for the molar mass of O_2.
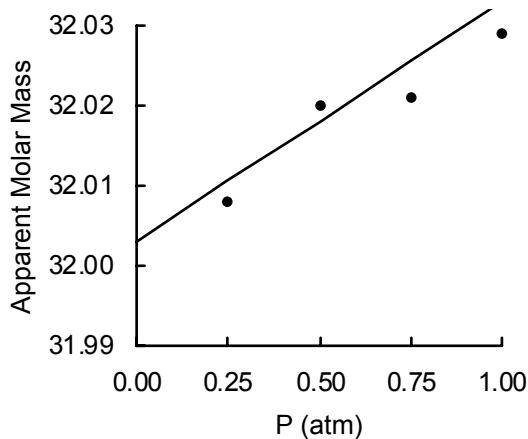
5.123   Use CO + $1/2O_2 \rightarrow CO_2$, instead of 2CO. First, find the moles of CO and O_2 by using the ideal gas law.

$$n_{CO} = \frac{PV}{RT} = \frac{(0.500 \text{ atm}) (2.00 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol}) (300 \text{ K})} = 0.040\underline{6}2 \text{ mol}$$

$$n_{O_2} = \frac{PV}{RT} = \frac{(1.00 \text{ atm}) (1.00 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol}) (300 \text{ K})} = 0.040\underline{6}2 \text{ mol}$$

There are equal amounts of CO and O_2, but (from the equation) only half as many moles of O_2 as CO are required for the reaction. Therefore, when 0.04062 moles of CO have been consumed, only 0.04062/2 moles of O_2 will have been used up. Then 0.04062/2 mol O_2 will remain, and 0.04062 mol CO_2 will have been produced. At the end,

$$n_{CO} = 0 \text{ mol}; \quad n_{O_2} = 0.0202 \text{ mol}; \quad \text{and } n_{CO_2} = 0.04062 \text{ mol}$$

However, the total volume with the valve open is 3.00 L, so the partial pressures of O_2 and CO_2 must be calculated from the ideal gas law for each:

$$\frac{nRT}{V} = \frac{(0.0203 \text{ mol } O_2)(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(300 \text{ K})}{3.00 \text{ L}}$$

$$= 0.16\underline{6}58 = 0.167 \text{ atm } O_2$$

$$\frac{nRT}{V} = \frac{(0.04062 \text{ mol } CO_2)(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(300 \text{ K})}{3.00 \text{ L}}$$

$$= 0.33\underline{3}32 = 0.333 \text{ atm } CO_2$$

5.124   Use $H_2 + 1/2O_2 \rightarrow H_2O$. First, find the moles of H_2 and O_2 by using the ideal gas law.

$$n_{H_2} = \frac{PV}{RT} = \frac{(0.500 \text{ atm})(2.00 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(300 \text{ K})} = 0.040\underline{6}2 \text{ mol}$$

$$n_{O_2} = \frac{PV}{RT} = \frac{(1.00 \text{ atm})(1.00 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(300 \text{ K})} = 0.040\underline{6}2 \text{ mol}$$

When all the H_2 has been consumed, half of the O_2 will remain. At the end, there will be 0.040$\underline{6}$2 mol of H_2O formed. However, the final volume is 3.00 L, so the partial pressures of H_2O and O_2 must each be calculated from the ideal gas law:

$$\frac{nRT}{V} = \frac{(0.04062 \text{ mol } H_2O)(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(300 \text{ K})}{3.00 \text{ L}}$$

$$= 0.33\underline{3}3 = 0.333 \text{ atm } H_2O$$

$$\frac{nRT}{V} = \frac{(0.02031 \text{ mol } O_2)(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(300 \text{ K})}{3.00 \text{ L}}$$

$$= 0.166\underline{6}6 = 0.167 \text{ atm } O_2$$

■ Solutions to Cumulative-Skills Problems

5.125   Assume a 100.0-g sample, giving 85.2 g CH_4 and 14.8 g C_2H_6. Convert each to moles:

$$85.2 \text{ g } CH_4 \times \frac{1 \text{ mol } CH_4}{16.04 \text{ g } CH_4} = 5.3\underline{1}1 \text{ mol } CH_4$$

$$14.8 \text{ g } C_2H_6 \times \frac{1 \text{ mol } CH_4}{16.04 \text{ g } CH_4} = 0.49\underline{2}1 \text{ mol } C_2H_6$$

$$V_{CH_4} = \frac{(5.311 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(291 \text{ K})}{(771/760) \text{ atm}} = 12\underline{5}.03 \text{ L}$$

$$V_{C_2H_6} = \frac{(0.4921 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(291 \text{ K})}{(771/760) \text{ atm}} = 11.\underline{5}8 \text{ L}$$

The density is calculated as follows:

$$d = \frac{85.2 \text{ g } CH_4 + 14.8 \text{ g } C_2H_6}{(125.03 + 11.58) \text{ L}} = 0.73\underline{1}9 = 0.732 \text{ g/L}$$

5.126 Assume a 100.0-g sample, giving 34.3 g He, 51.7 g N_2, and 14.0 g O_2. Convert to moles:

$$34.3 \text{ g He} \times \frac{1 \text{ mol He}}{4.00 \text{ g}} = 8.5\underline{7}50 \text{ mol He}$$

$$51.7 \text{ g } N_2 \times \frac{1 \text{ mol } N_2}{28.02 \text{ g}} = 1.8\underline{4}51 \text{ mol } N_2$$

$$14.0 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} = 0.43\underline{7}50 \text{ mol } O_2$$

$$V_{He} = \frac{(8.575 \text{ mol })(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(295 \text{ K})}{(755/760) \text{ atm}} = 208.\underline{9} \text{ L}$$

$$V_{N_2} = \frac{(1.8451 \text{ mol })(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(295 \text{ K})}{(755/760) \text{ atm}} = 44.\underline{9}6 \text{ L}$$

$$V_{O_2} = \frac{(0.4375 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(295 \text{ K})}{(755/760) \text{ atm}} = 10.\underline{6}6 \text{ L}$$

The density is calculated as follows:

$$d = \frac{34.3 \text{ g He} + 51.7 \text{ g } N_2 + 14.0 \text{ g } O_2}{(208.9 + 44.96 + 10.66) \text{ L}} = 0.37\underline{8}02 = 0.378 \text{ g/L}$$

5.127 First, subtract the height of mercury equivalent to the 25.00 cm (250 mm) of water inside the tube from 771 mmHg to get P_{gas}. Then, subtract the vapor pressure of water, 18.7 mmHg, from P_{gas} to get P_{O_2}

$$h_{Hg} = \frac{(h_W)(d_W)}{d_{Hg}} = \frac{250 \text{ mm} \times 0.99987 \text{ g/cm}^3}{13.596 \text{ g/cm}^3} = 18.\underline{3}8 \text{ mmHg}$$

$$P_{gas} = P - P_{25 \text{ cm water}} = 771 \text{ mmHg} - 18.38 \text{ mmHg} = 75\underline{2}.62 \text{ mmHg}$$

$$P_{O_2} = 75\underline{2}.62 \text{ mmHg} - 18.7 \text{ mmHg} = 73\underline{3}.92 \text{ mmHg}$$

$$n = \frac{PV}{RT} = \frac{(733.92/760 \text{ atm}) (0.0310 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol}) (294 \text{ K})} = 0.00124\underline{0}8 \text{ mol } O_2$$

$$\text{Mass} = (2 \times 0.0012408) \text{ mol } Na_2O_2 \times \frac{77.98 \text{ g } Na_2O_2}{1 \text{ mol } Na_2O_2}$$

$$= 0.19\underline{3}5 = 0.194 \text{ g } Na_2O_2$$

5.128 Proceed as in the previous problem, subtracting the height of water and then vapor pressure to find P_{H_2} .

$$h_{Hg} = \frac{(h_W)(d_W)}{d_{Hg}} = \frac{310 \text{ mm} \times 0.99987 \text{ g/cm}^3}{13.596 \text{ g/cm}^3} = 22.\underline{8}0 \text{ mmHg}$$

$$P_{gas} = P - P_{31 \text{ cm water}} = 751 \text{ mmHg} - 22.80 \text{ mmHg} = 72\underline{8}.2 \text{ mmHg}$$

$$P_{H_2} = (728.2 - 15.5) \text{ mmHg} = 71\underline{2}.7 \text{ mmHg}$$

$$n = \frac{PV}{RT} = \frac{(712.7 / 760 \text{ atm}) (0.0221 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(290 \text{ K})} = 0.00087\underline{0}87 \text{ mol } H_2$$

$$\text{Mass} = 0.00087\underline{0}87 \text{ mol } H_2 \times \frac{1 \text{ mol Zn}}{1 \text{ mol } H_2} \times \frac{65.39 \text{ g Zn}}{1 \text{ mol Zn}}$$

$$= 0.056\underline{9}46 = 0.0569 \text{ g Zn}$$

5.129 First find the moles of CO_2:

$$n = \frac{PV}{RT} = \frac{(785/760 \text{ atm}) (1.94 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol}) (298 \text{ K})} = 0.081\underline{9}4 \text{ mol CO}$$

Set up one equation in one unknown: x = mol $CaCO_3$; (0.08194 - x) = mol $MgCO_3$.

7.85 g = (100.1 g/mol)x + (84.32 g/mol)(0.08194 - x)

$$x = \frac{(7.85 - 6.9092)}{(100.1 - 84.32)} = 0.05\underline{9}62 \text{ mol } CaCO_3$$

(0.08194 - x) = 0.02\underline{2}32 mol $MgCO_3$

$$CaCO_3 = \frac{0.05962 \text{ mol } CaCO_3 \times 100.1 \text{ g/mol}}{7.85 \text{ g}} \times 100\% = 7\underline{6}.02 = 76\%$$

Percent $MgCO_3$ = 100.00% - 76.02% = 2\underline{3}.98 = 24%

5.130   First find the moles of H_2S:

$$n = \frac{PV}{RT} = \frac{(745/760 \text{ atm})(1.049 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(296 \text{ K})} = 0.042\underline{3}3 \text{ mol } H_2S$$

Set up one equation in one unknown: x = mol ZnS; (0.04233 - x) = mol PbS.

6.12 g = (97.46 g/mol)x + (239.25 g/mol)(0.04233 - x)

$$x = \frac{(6.12 - 10.127)}{(97.46 - 239.25)} = 0.028\underline{2}6 \text{ mol } ZnS$$

(0.04233 - x) = 0.0104\underline{0}70 mol PbS

$$\text{Percent ZnS} = \frac{0.02826 \text{ mol ZnS} \times 97.46 \text{ g/mol}}{6.12 \text{ g}} \times 100\% = 45.\underline{0}0 = 45.0\%$$

Percent PbS = 55.0%

5.131   Write a mass-balance equation to solve for the moles of each gas using 28.01 g/mol for the molar mass of N_2 and 20.18 g/mol for the molar mass of Ne. Let y equal mol of each gas:

28.01 y + 20.18 y = 10.0 g

y = (10.0 ÷ 48.19) = 0.207\underline{5}1 mol

Total moles (n) = 0.41502 mol (use below)

Use the ideal gas law to calculate the volume, which can then be used to calculate density.

$$\frac{(0.41502 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(500 \text{ K})}{15.00 \text{ atm}} = 1.1\underline{3}52$$

d = 10.0 g ÷ 1.1352 L = 8.8\underline{0}8 = 8.81 g/L

5.132 Write a mass-balance equation to solve for moles of each using 39.95 g/mol for the molar mass of Ar and 20.18 g/mol for the molar mass of Ne. Let y equal mol of Ar and 2y equal mol of Ne:

$$39.95\,y + 20.18\,(2y) = 50.0\,g$$

$$y = (50.0 \div 80.31) = 0.62\underline{2}5\,mol$$

Total moles (n) = mol Ar + mol Ne = 0.6225 + (2)(0.6225) = 1.86\underline{7}7 mol (use below)

The volume of the mixture = 50.0 g ÷ 4.00 g/L = 12.5 L

Use the ideal gas law to calculate the total pressure

$$P(total) = \frac{(1.8677\,mol)(0.08206\,L \bullet atm/K \bullet mol)(350\,K)}{12.5\,L} = 4.2\underline{9}2\,atm$$

The partial pressure of Ne = 2/3 P(total) = 2/3(4.292 atm) = 2.8\underline{6}1 = 2.86 atm.

5.133 Rearrange the equation $PM_m = dRT$ to find the quantity RT/M_m. Then, plug into the equation for the root-mean-square speed.

$$\frac{RT}{M_m} = \frac{P}{d}$$

$$u = \sqrt{\frac{3RT}{M_m}} = \sqrt{\frac{3P}{d}} = \sqrt{\frac{3\left(\frac{675}{760}\right)atm \times (8.31\,kg \bullet m^2/s^2 \bullet K \bullet mol)}{(3.00 \times 10^{-3}\,kg/L)(0.08206\,L \bullet atm/K \bullet mol)}}$$

$$= 2.9\underline{9}90 \times 10^2 = 3.00 \times 10^2\,m/s$$

5.134 Rearrange the root-mean-square speed equation to get $M_m = 3RT/u^2$.

$$M_m = \frac{(3)(8.31\,kg \bullet m^2/(s^2 \bullet mol \bullet K)(298.2\,K)}{(5.00 \times 10^2\,m/s)^2} = 0.02973\underline{7}\,kg/mol$$

$$= 29.\underline{7}4\,g/mol$$

Now, rearrange the equation $PM_m = dRT$ to find the density, $d = PM_m/RT$.

$$d = \frac{PM_m}{RT} = \frac{(2.50\,atm)(29.74\,g/mol)}{(0.08206\,L \bullet atm/K \bullet mol)(298.2\,K)} = 3.0\underline{3}8 = 3.04\,g/L$$

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