CHAPTER 5
GASES

5.13 $562 \text{ mmHg} \times \dfrac{1 \text{ atm}}{760 \text{ mmHg}} = \mathbf{0.739 \text{ atm}}$

5.14 **Strategy:** Because 1 atm = 760 mmHg, the following conversion factor is needed to obtain the pressure in atmospheres.

$$\frac{1 \text{ atm}}{760 \text{ mmHg}}$$

For the second conversion, 1 atm = 101.325 kPa.

Solution:

$$\textbf{? atm} = 606 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = \mathbf{0.797 \text{ atm}}$$

$$\textbf{? kPa} = 0.797 \text{ atm} \times \frac{101.325 \text{ kPa}}{1 \text{ atm}} = \mathbf{80.8 \text{ kPa}}$$

5.17 **(a)** If the final temperature of the sample is above the boiling point, it would still be in the gas phase. The diagram that best represents this is choice **(d)**.

(b) If the final temperature of the sample is below its boiling point, it will condense to a liquid. The liquid will have a vapor pressure, so some of the sample will remain in the gas phase. The diagram that best represents this is choice **(b)**.

5.18 **(1)** Recall that $V \propto \dfrac{1}{P}$. As the pressure is tripled, the volume will decrease to $\frac{1}{3}$ of its original volume, assuming constant n and T. The correct choice is **(b)**.

(2) Recall that $V \propto T$. As the temperature is doubled, the volume will also double, assuming constant n and P. The correct choice is **(a)**. The depth of color indicates the density of the gas. As the volume increases at constant moles of gas, the density of the gas will decrease. This decrease in gas density is indicated by the lighter shading.

(3) Recall that $V \propto n$. Starting with n moles of gas, adding another n moles of gas ($2n$ total) will double the volume. The correct choice is **(c)**. The density of the gas will remain the same as moles are doubled and volume is doubled.

(4) Recall that $V \propto T$ and $V \propto \dfrac{1}{P}$. Halving the temperature would decrease the volume to $\frac{1}{2}$ its original volume. However, reducing the pressure to $\frac{1}{4}$ its original value would increase the volume by a factor of 4. Combining the two changes, we have

$$\frac{1}{2} \times 4 = 2$$

The volume will double. The correct choice is **(a)**.

5.19 $P_1 = 0.970 \text{ atm}$ $P_2 = 0.541 \text{ atm}$

$V_1 = 725 \text{ mL}$ $\boldsymbol{V_2 = ?}$

$$P_1V_1 = P_2V_2$$

$$\boldsymbol{V_2} = \frac{P_1V_1}{P_2} = \frac{(0.970 \text{ atm})(725 \text{ mL})}{0.541 \text{ atm}} = \mathbf{1.30 \times 10^3\ mL}$$

5.20 Temperature and amount of gas do not change in this problem ($T_1 = T_2$ and $n_1 = n_2$). Pressure and volume change; it is a Boyle's law problem.

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

$$P_1V_1 = P_2V_2$$

$V_2 = 0.10\ V_1$

$$P_2 = \frac{P_1V_1}{V_2}$$

$$\boldsymbol{P_2} = \frac{(5.3 \text{ atm})V_1}{0.10V_1} = \mathbf{53\ atm}$$

5.21 $P_1 = 1.00 \text{ atm} = 760 \text{ mmHg}$ $\boldsymbol{P_2 = ?}$

$V_1 = 5.80 \text{ L}$ $V_2 = 9.65 \text{ L}$

$$P_1V_1 = P_2V_2$$

$$\boldsymbol{P_2} = \frac{P_1V_1}{V_2} = \frac{(760 \text{ mmHg})(5.80 \text{ L})}{9.65 \text{ L}} = \mathbf{457\ mmHg}$$

5.22 **(a)**

Strategy: The amount of gas and its temperature remain constant, but both the pressure and the volume change. What equation would you use to solve for the final volume?

Solution: We start with Equation (5.9) of the text.

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

Because $n_1 = n_2$ and $T_1 = T_2$,

$$P_1V_1 = P_2V_2$$

which is Boyle's Law. The given information is tabulated below.

Initial conditions	Final Conditions
$P_1 = 1.2 \text{ atm}$	$P_2 = 6.6 \text{ atm}$
$V_1 = 3.8 \text{ L}$	$V_2 = ?$

The final volume is given by:

$$V_2 = \frac{P_1V_1}{P_2}$$

$$\boldsymbol{V_2} = \frac{(1.2\ \text{atm})(3.8\ \text{L})}{(6.6\ \text{atm})} = \mathbf{0.69\ L}$$

Check: When the pressure applied to the sample of air is increased from 1.2 atm to 6.6 atm, the volume occupied by the sample will decrease. Pressure and volume are inversely proportional. The final volume calculated is less than the initial volume, so the answer seems reasonable.

(b)
Strategy: The amount of gas and its temperature remain constant, but both the pressure and the volume change. What equation would you use to solve for the final pressure?

Solution: You should also come up with the equation $P_1V_1 = P_2V_2$ for this problem. The given information is tabulated below.

Initial conditions	Final Conditions
$P_1 = 1.2$ atm	$P_2 = ?$
$V_1 = 3.8$ L	$V_2 = 0.075$ L

The final pressure is given by:

$$P_2 = \frac{P_1V_1}{V_2}$$

$$\boldsymbol{P_2} = \frac{(1.2\ \text{atm})(3.8\ \text{L})}{(0.075\ \text{L})} = \mathbf{61\ atm}$$

Check: To decrease the volume of the gas fairly dramatically from 3.8 L to 0.075 L, the pressure must be increased substantially. A final pressure of 61 atm seems reasonable.

5.23 $T_1 = 25° + 273° = 298\ \text{K}$ $T_2 = 88° + 273° = 361\ \text{K}$

$V_1 = 36.4\ \text{L}$ $\boldsymbol{V_2 = ?}$

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

$$\boldsymbol{V_2} = \frac{V_1T_2}{T_1} = \frac{(36.4\ \text{L})(361\ \text{K})}{298\ \text{K}} = \mathbf{44.1\ L}$$

5.24 **Strategy:** The amount of gas and its pressure remain constant, but both the temperature and the volume change. What equation would you use to solve for the final temperature? What temperature unit should we use?

Solution: We start with Equation (5.9) of the text.

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

Because $n_1 = n_2$ and $P_1 = P_2$,

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

which is Charles' Law. The given information is tabulated below.

Initial conditions	Final Conditions
$T_1 = (88 + 273)\text{K} = 361\text{ K}$	$T_2 = ?$
$V_1 = 9.6\text{ L}$	$V_2 = 3.4\text{ L}$

The final temperature is given by:

$$T_2 = \frac{T_1V_2}{V_1}$$

$$\boldsymbol{T_2} = \frac{(361\text{ K})(3.4\text{ L})}{(9.6\text{ L})} = \mathbf{1.3 \times 10^2\ K}$$

5.25 The balanced equation is: $4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(g)$

Recall that Avogadro's Law states that the volume of a gas is directly proportional to the number of moles of gas at constant temperature and pressure. The ammonia and nitric oxide coefficients in the balanced equation are the same, so **one volume** of nitric oxide must be obtained from **one volume** of ammonia.

Could you have reached the same conclusion if you had noticed that nitric oxide is the only nitrogen-containing product and that ammonia is the only nitrogen-containing reactant?

5.26 This is a gas stoichiometry problem that requires knowledge of Avogadro's law to solve. Avogadro's law states that the volume of a gas is directly proportional to the number of moles of gas at constant temperature and pressure.

The volume ratio, 1 vol. Cl_2 : 3 vol. F_2 : 2 vol. product, can be written as a mole ratio, 1 mol Cl_2 : 3 mol F_2 : 2 mol product.

Attempt to write a balanced chemical equation. The subscript of F in the product will be three times the Cl subscript, because there are three times as many F atoms reacted as Cl atoms.

$$\mathbf{1}Cl_2(g) + \mathbf{3}F_2(g) \longrightarrow \mathbf{2}Cl_xF_{3x}(g)$$

Balance the equation. The x must equal one so that there are two Cl atoms on each side of the equation. If $x = 1$, the subscript on F is 3.

$$Cl_2(g) + 3F_2(g) \longrightarrow 2ClF_3(g)$$

The formula of the product is $\mathbf{ClF_3}$.

5.31 $$\boldsymbol{n} = \frac{PV}{RT} = \frac{(4.7\text{ atm})(2.3\text{ L})}{\left(0.0821\,\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273 + 32)\text{K}} = \mathbf{0.43\ mol}$$

5.32 **Strategy:** This problem gives the amount, volume, and temperature of CO gas. Is the gas undergoing a change in any of its properties? What equation should we use to solve for the pressure? What temperature unit should be used?

Solution: Because no changes in gas properties occur, we can use the ideal gas equation to calculate the pressure. Rearranging Equation (5.8) of the text, we write:

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

$$\boldsymbol{P} = \frac{(6.9\text{ mol})\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(62 + 273)\text{K}}{30.4\text{ L}} = \mathbf{6.2\ atm}$$

5.33 We solve the ideal gas equation for V.

$$\boldsymbol{V} = \frac{nRT}{P} = \frac{(5.6\text{ mol})\left(0.0821\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(128+273)\text{K}}{9.4\text{ atm}} = \mathbf{2.0 \times 10^1\ L}$$

5.34 In this problem, the moles of gas and the volume the gas occupies are constant ($V_1 = V_2$ and $n_1 = n_2$). Temperature and pressure change.

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

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

The given information is tabulated below.

Initial conditions	Final Conditions
$T_1 = (25 + 273)\text{K} = 298\text{ K}$	$T_2 = ?$
$P_1 = 0.800\text{ atm}$	$P_2 = 2.00\text{ atm}$

The final temperature is given by:

$$T_2 = \frac{T_1P_2}{P_1}$$

$$\boldsymbol{T_2} = \frac{(298\text{ K})(2.00\text{ atm})}{(0.800\text{ atm})} = \mathbf{745\ K = 472^\circ C}$$

5.35 Initial Conditions / Final Conditions

Initial Conditions	Final Conditions
$P_1 = 1.2\text{ atm}$	$P_2 = 3.00 \times 10^{-3}\text{ atm}$
$V_1 = 2.50\text{ L}$	$\boldsymbol{V_2 = ?}$
$T_1 = (25 + 273)\text{K} = 298\text{ K}$	$T_2 = (-23 + 273)\text{K} = 250\text{ K}$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\boldsymbol{V_2} = \frac{P_1V_1T_2}{T_1P_2} = \frac{(1.2\text{ atm})(2.50\text{ L})(250\text{ K})}{(298\text{ K})(3.00\times10^{-3}\text{ atm})} = \mathbf{8.4\times10^2\ L}$$

5.36 In this problem, the moles of gas and the volume the gas occupies are constant ($V_1 = V_2$ and $n_1 = n_2$). Temperature and pressure change.

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

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

The given information is tabulated below.

Initial conditions	Final Conditions
$T_1 = 273\text{ K}$	$T_2 = (250 + 273)\text{K} = 523\text{ K}$
$P_1 = 1.0\text{ atm}$	$P_2 = ?$

The final pressure is given by:

$$P_2 = \frac{P_1T_2}{T_1}$$

$$\boldsymbol{P_2} = \frac{(1.0\text{ atm})(523\text{ K})}{273\text{ K}} = \mathbf{1.9\ atm}$$

5.37 Note that the statement "...its absolute temperature is decreased by one-half" implies that $\frac{T_2}{T_1} = 0.50$.

Similarly, the statement "...pressure is decreased to one-third of its original pressure" indicates $\frac{P_2}{P_1} = \frac{1}{3}$.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\boldsymbol{V_2} = V_1\left(\frac{P_1}{P_2}\right)\left(\frac{T_2}{T_1}\right) = 6.0\text{ L}\times3.0\times0.50 = \mathbf{9.0\ L}$$

5.38 In this problem, the moles of gas and the pressure on the gas are constant ($n_1 = n_2$ and $P_1 = P_2$). Temperature and volume are changing.

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

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

The given information is tabulated below.

Initial conditions	Final Conditions
$T_1 = (20.1 + 273)\text{ K} = 293.1\text{ K}$	$T_2 = (36.5 + 273)\text{K} = 309.5\text{ K}$
$V_1 = 0.78\text{ L}$	$V_2 = ?$

The final volume is given by:

$$V_2 = \frac{V_1T_2}{T_1}$$

$$\mathbf{V_2} = \frac{(0.78\text{ L})(309.5\text{ K})}{(293.1\text{ K})} = \mathbf{0.82\ L}$$

5.39 $$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\mathbf{V_1} = \frac{P_2V_2T_1}{P_1T_2} = \frac{(0.60\text{ atm})(94\text{ mL})(66 + 273)\text{K}}{(0.85\text{ atm})(45 + 273)\text{K}} = \mathbf{71\ mL}$$

5.40 In the problem, temperature and pressure are given. If we can determine the moles of CO_2, we can calculate the volume it occupies using the ideal gas equation.

$$?\text{ mol CO}_2 = 88.4\text{ g CO}_2 \times \frac{1\text{ mol CO}_2}{44.01\text{ g CO}_2} = 2.01\text{ mol CO}_2$$

We now substitute into the ideal gas equation to calculate volume of CO_2.

$$\mathbf{V_{CO_2}} = \frac{nRT}{P} = \frac{(2.01\text{ mol})\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273\text{ K})}{(1\text{ atm})} = \mathbf{45.1\ L}$$

Alternatively, we could use the fact that 1 mole of an ideal gas occupies a volume of 22.41 L at STP. After calculating the moles of CO_2, we can use this fact as a conversion factor to convert to volume of CO_2.

$$?\text{ L CO}_2 = 2.01\text{ mol CO}_2 \times \frac{22.41\text{ L}}{1\text{ mol}} = \mathbf{45.0\ L\ CO_2}$$

The slight difference in the results of our two calculations is due to rounding the volume occupied by 1 mole of an ideal gas to 22.41 L.

5.41 $$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\mathbf{V_2} = \frac{P_1V_1T_2}{P_2T_1} = \frac{(772\text{ mmHg})(6.85\text{ L})(273\text{ K})}{(760\text{ mmHg})(35 + 273)\text{K}} = \mathbf{6.17\ L}$$

5.42 The molar mass of CO_2 = 44.01 g/mol. Since $PV = nRT$, we write:

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

$$\boldsymbol{P} = \frac{\left(0.050\ \text{g} \times \frac{1\ \text{mol}}{44.01\ \text{g}}\right)\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(30 + 273)\text{K}}{4.6\ \text{L}} = \mathbf{6.1 \times 10^{-3}\ atm}$$

5.43 Solve for the number of moles of gas using the ideal gas equation.

$$n = \frac{PV}{RT} = \frac{(1.00\ \text{atm})(0.280\ \text{L})}{\left(0.0821\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273\ \text{K})} = 0.0125\ \text{mol}$$

Solving for the molar mass:

$$\mathcal{M} = \frac{\text{mass (in g)}}{\text{mol}} = \frac{0.400\ \text{g}}{0.0125\ \text{mol}} = \mathbf{32.0\ g/mol}$$

5.44 **Strategy:** We can calculate the molar mass of a gas if we know its density, temperature, and pressure. What temperature and pressure units should we use?

Solution: We need to use Equation (5.12) of the text to calculate the molar mass of the gas.

$$\mathcal{M} = \frac{dRT}{P}$$

Before substituting into the above equation, we need to calculate the density and check that the other known quantities (P and T) have the appropriate units.

$$d = \frac{7.10\ \text{g}}{5.40\ \text{L}} = 1.31\ \text{g/L}$$

$$T = 44° + 273° = 317\ \text{K}$$

$$P = 741\ \text{torr} \times \frac{1\ \text{atm}}{760\ \text{torr}} = 0.975\ \text{atm}$$

Calculate the molar mass by substituting in the known quantities.

$$\mathcal{M} = \frac{\left(1.31\ \frac{\text{g}}{\text{L}}\right)\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(317\ \text{K})}{0.975\ \text{atm}} = \mathbf{35.0\ g/mol}$$

Alternatively, we can solve for the molar mass by writing:

$$\text{molar mass of compound} = \frac{\text{mass of compound}}{\text{moles of compound}}$$

Mass of compound is given in the problem (7.10 g), so we need to solve for moles of compound in order to calculate the molar mass.

$$n = \frac{PV}{RT}$$

$$n = \frac{(0.975\text{ atm})(5.40\text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(317\text{ K})} = 0.202\text{ mol}$$

Now, we can calculate the molar mass of the gas.

$$\textbf{molar mass of compound} = \frac{\text{mass of compound}}{\text{moles of compound}} = \frac{7.10\text{ g}}{0.202\text{ mol}} = \textbf{35.1 g/mol}$$

5.45 First calculate the moles of ozone (O_3) using the ideal gas equation.

$$n = \frac{PV}{RT} = \frac{(1.0\times10^{-3}\text{ atm})(1.0\text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(250\text{ K})} = 4.9\times10^{-5}\text{ mol O}_3$$

Use Avogadro's number as a conversion factor to convert to molecules of O_3.

$$\textbf{molecules O}_\mathbf{3} = (4.9\times10^{-5}\text{ mol O}_3)\times\frac{6.022\times10^{23}\text{ O}_3\text{ molecules}}{1\text{ mol O}_3} = \mathbf{3.0\times10^{19}\ O_3\ molecules}$$

5.46 The number of particles in 1 L of gas at STP is:

$$\text{Number of particles} = 1.0\text{ L}\times\frac{1\text{ mol}}{22.414\text{ L}}\times\frac{6.022\times10^{23}\text{ particles}}{1\text{ mol}} = 2.7\times10^{22}\text{ particles}$$

$$\textbf{Number of N}_\mathbf{2}\textbf{ molecules} = \left(\frac{78\%}{100\%}\right)(2.7\times10^{22}\text{ particles}) = \mathbf{2.1\times10^{22}\ N_2\ molecules}$$

$$\textbf{Number of O}_\mathbf{2}\textbf{ molecules} = \left(\frac{21\%}{100\%}\right)(2.7\times10^{22}\text{ particles}) = \mathbf{5.7\times10^{21}\ O_2\ molecules}$$

$$\textbf{Number of Ar atoms} = \left(\frac{1\%}{100\%}\right)(2.7\times10^{22}\text{ particles}) = \mathbf{3\times10^{20}\ Ar\ atoms}$$

5.47 The density is given by:

$$\textbf{density} = \frac{\text{mass}}{\text{volume}} = \frac{4.65\text{ g}}{2.10\text{ L}} = \textbf{2.21 g/L}$$

Solving for the molar mass:

$$\textbf{molar mass} = \frac{dRT}{P} = \frac{(2.21\text{ g/L})\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(27+273)\text{K}}{(1.00\text{ atm})} = \textbf{54.4 g/mol}$$

5.48 The density can be calculated from the ideal gas equation.

$$d = \frac{P\mathcal{M}}{RT}$$

$\mathcal{M} = 1.008\ \text{g/mol} + 79.90\ \text{g/mol} = 80.91\ \text{g/mol}$
$T = 46° + 273° = 319\ \text{K}$
$P = 733\ \text{mmHg} \times \dfrac{1\ \text{atm}}{760\ \text{mmHg}} = 0.964\ \text{atm}$

$$\boldsymbol{d} = \frac{(0.964\ \text{atm})\left(\dfrac{80.91\ \text{g}}{1\ \text{mol}}\right)}{319\ \text{K}} \times \frac{\text{mol}\cdot\text{K}}{0.0821\ \text{L}\cdot\text{atm}} = \mathbf{2.98\ g/L}$$

Alternatively, we can solve for the density by writing:

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

Assuming that we have 1 mole of HBr, the mass is 80.91 g. The volume of the gas can be calculated using the ideal gas equation.

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

$$V = \frac{(1\ \text{mol})\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(319\ \text{K})}{0.964\ \text{atm}} = 27.2\ \text{L}$$

Now, we can calculate the density of HBr gas.

$$\textbf{density} = \frac{\text{mass}}{\text{volume}} = \frac{80.91\ \text{g}}{27.2\ \text{L}} = \mathbf{2.97\ g/L}$$

5.49 **METHOD 1:**

The empirical formula can be calculated from mass percent data. The molar mass can be calculated using the ideal gas equation. The molecular formula can then be determined.

To calculate the empirical formula, assume 100 g of substance.

$$64.9\ \text{g C} \times \frac{1\ \text{mol C}}{12.01\ \text{g C}} = 5.40\ \text{mol C}$$

$$13.5\ \text{g H} \times \frac{1\ \text{mol H}}{1.008\ \text{g H}} = 13.4\ \text{mol H}$$

$$21.6\ \text{g O} \times \frac{1\ \text{mol O}}{16.00\ \text{g O}} = 1.35\ \text{mol O}$$

This gives the formula $C_{5.40}H_{13.4}O_{1.35}$. Dividing by 1.35 gives the empirical formula, $\mathbf{C_4H_{10}O}$.

To calculate the molar mass, first calculate the number of moles of gas using the ideal gas equation.

$$n = \frac{PV}{RT} = \frac{\left(750\text{ mmHg} \times \dfrac{1\text{ atm}}{760\text{ mmHg}}\right)(1.00\text{ L})}{\left(0.0821\,\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(120 + 273)\text{K}} = 0.0306\text{ mol}$$

Solving for the molar mass:

$$\mathcal{M} = \frac{\text{mass (in g)}}{\text{mol}} = \frac{2.30\text{ g}}{0.0306\text{ mol}} = \mathbf{75.2\ g/mol}$$

The empirical mass is 74.0 g/mol which is essentially the same as the molar mass. In this case, the molecular formula is the same as the empirical formula, $\mathbf{C_4H_{10}O}$.

METHOD 2:

First calculate the molar mass using the ideal gas equation.

$$n = \frac{PV}{RT} = \frac{\left(750\text{ mmHg} \times \dfrac{1\text{ atm}}{760\text{ mmHg}}\right)(1.00\text{ L})}{\left(0.0821\,\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(120 + 273)\text{K}} = 0.0306\text{ mol}$$

Solving for the molar mass:

$$\mathcal{M} = \frac{\text{mass (in g)}}{\text{mol}} = \frac{2.30\text{ g}}{0.0306\text{ mol}} = \mathbf{75.2\ g/mol}$$

Next, multiply the mass % (converted to a decimal) of each element by the molar mass to convert to grams of each element. Then, use the molar mass to convert to moles of each element.

$$\boldsymbol{n}_{\mathbf{C}} = (0.649) \times (75.2\text{ g}) \times \frac{1\text{ mol C}}{12.01\text{ g C}} = \mathbf{4.06\ mol\ C}$$

$$\boldsymbol{n}_{\mathbf{H}} = (0.135) \times (75.2\text{ g}) \times \frac{1\text{ mol H}}{1.008\text{ g H}} = \mathbf{10.07\ mol\ H}$$

$$\boldsymbol{n}_{\mathbf{O}} = (0.216) \times (75.2\text{ g}) \times \frac{1\text{ mol O}}{16.00\text{ g O}} = \mathbf{1.02\ mol\ O}$$

Since we used the molar mass to calculate the moles of each element present in the compound, this method directly gives the molecular formula. The formula is $\mathbf{C_4H_{10}O}$.

5.50 This is an extension of an ideal gas law calculation involving molar mass. If you determine the molar mass of the gas, you will be able to determine the molecular formula from the empirical formula.

$$\mathcal{M} = \frac{dRT}{P}$$

Calculate the density, then substitute its value into the equation above.

$$d = \frac{0.100\text{ g}}{22.1\text{ mL}} \times \frac{1000\text{ mL}}{1\text{ L}} = 4.52\text{ g/L}$$

$$T(\text{K}) = 20° + 273° = 293\ \text{K}$$

$$\mathcal{M} = \frac{\left(4.52\ \frac{\text{g}}{\text{L}}\right)\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(293\ \text{K})}{1.02\ \text{atm}} = 107\ \text{g/mol}$$

Compare the empirical mass to the molar mass.

$$\text{empirical mass} = 32.07\ \text{g/mol} + 4(19.00\ \text{g/mol}) = 108.07\ \text{g/mol}$$

Remember, the molar mass will be a whole number multiple of the empirical mass. In this case, the $\frac{\text{molar mass}}{\text{empirical mass}} \approx 1$. Therefore, the molecular formula is the same as the empirical formula, $\mathbf{SF_4}$.

5.51 In addition to a mole ratio, the coefficients from a balanced equation can represent the volume ratio in which the gases in the equation react and are produced. Recall that Avogadro's Law states that $V \propto n$. See Figure 5.10 of the text. We can use this volume ratio to convert from liters of NO to liters of NO_2.

$$9.0\ \text{L NO} \times \frac{2\ \text{volumes NO}_2}{2\ \text{volumes NO}} = \mathbf{9.0\ L\ NO_2}$$

5.52 **Strategy:** From the moles of CH_4 reacted, we can calculate the moles of CO_2 produced. From the balanced equation, we see that 1 mol $CH_4 \simeq$ 1 mol CO_2. Once moles of CO_2 are determined, we can use the ideal gas equation to calculate the volume of CO_2.

Solution: First let's calculate moles of CO_2 produced.

$$?\ \text{mol CO}_2 = 15.0\ \text{mol CH}_4 \times \frac{1\ \text{mol CO}_2}{1\ \text{mol CH}_4} = 15.0\ \text{mol CO}_2$$

Now, we can substitute moles, temperature, and pressure into the ideal gas equation to solve for volume of CO_2.

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

$$V_{\mathbf{CO_2}} = \frac{(15.0\ \text{mol})\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(23 + 273)\text{K}}{0.985\ \text{atm}} = \mathbf{3.70 \times 10^2\ L}$$

5.53 If we can calculate the moles of S, we can use the mole ratio from the balanced equation to calculate the moles of SO_2. Once we know the moles of SO_2, we can determine the volume of SO_2 using the ideal gas equation.

$$(2.54 \times 10^3\ \text{g S}) \times \frac{1\ \text{mol S}}{32.07\ \text{g S}} \times \frac{1\ \text{mol SO}_2}{1\ \text{mol S}} = 79.2\ \text{mol SO}_2$$

$$\boldsymbol{V} = \frac{nRT}{P} = \frac{(79.2\ \text{mol})\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(303.5\ \text{K})}{1.12\ \text{atm}} = 1.76 \times 10^3\ \text{L} = \mathbf{1.76 \times 10^6\ mL\ SO_2}$$

5.54 From the amount of glucose reacted (5.97 g), we can calculate the theoretical yield of CO_2. We can then compare the theoretical yield to the actual yield given in the problem (1.44 L) to determine the percent yield.

First, let's determine the moles of CO_2 that can be produced theoretically. Then, we can use the ideal gas equation to determine the volume of CO_2.

$$? \text{ mol CO}_2 = 5.97 \text{ g glucose} \times \frac{1 \text{ mol glucose}}{180.2 \text{ g glucose}} \times \frac{2 \text{ mol CO}_2}{1 \text{ mol glucose}} = 0.0663 \text{ mol CO}_2$$

Now, substitute moles, pressure, and temperature into the ideal gas equation to calculate the volume of CO_2.

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

$$V_{\text{CO}_2} = \frac{(0.0663 \text{ mol})\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(293 \text{ K})}{0.984 \text{ atm}} = 1.62 \text{ L}$$

This is the theoretical yield of CO_2. The actual yield, which is given in the problem, is 1.44 L. We can now calculate the percent yield.

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$\textbf{percent yield} = \frac{1.44 \text{ L}}{1.62 \text{ L}} \times 100\% = \mathbf{88.9\%}$$

5.55 If you determine the molar mass of the gas, you will be able to determine the molecular formula from the empirical formula. First, let's calculate the molar mass of the compound.

$$n = \frac{PV}{RT} = \frac{\left(97.3 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}\right)(0.378 \text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(77 + 273)\text{K}} = 0.00168 \text{ mol}$$

Solving for the molar mass:

$$\mathcal{M} = \frac{\text{mass (in g)}}{\text{mol}} = \frac{0.2324 \text{ g}}{0.00168 \text{ mol}} = \textbf{138 g/mol}$$

To calculate the empirical formula, first we need to find the mass of F in 0.2631 g of CaF_2.

$$0.2631 \text{ g CaF}_2 \times \frac{1 \text{ mol CaF}_2}{78.08 \text{ g CaF}_2} \times \frac{2 \text{ mol F}}{1 \text{ mol CaF}_2} \times \frac{19.00 \text{ g F}}{1 \text{ mol F}} = 0.1280 \text{ g F}$$

Since the compound only contains P and F, the mass of P in the 0.2324 g sample is:

$$0.2324 \text{ g} - 0.1280 \text{ g} = 0.1044 \text{ g P}$$

Now, we can convert masses of P and F to moles of each substance.

$$? \text{ mol P} = 0.1044 \text{ g P} \times \frac{1 \text{ mol P}}{30.97 \text{ g P}} = 0.003371 \text{ mol P}$$

$$? \text{ mol F} = 0.1280 \text{ g F} \times \frac{1 \text{ mol F}}{19.00 \text{ g F}} = 0.006737 \text{ mol F}$$

Thus, we arrive at the formula $P_{0.003371}F_{0.006737}$. Dividing by the smallest number of moles (0.003371 mole) gives the empirical formula PF_2.

To determine the molecular formula, divide the molar mass by the empirical mass.

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{138 \text{ g}}{68.97 \text{ g}} \approx 2$$

Hence, the molecular formula is $(PF_2)_2$ or $\mathbf{P_2F_4}$.

5.56 **Strategy:** We can calculate the moles of M reacted, and the moles of H_2 gas produced. By comparing the number of moles of M reacted to the number of moles H_2 produced, we can determine the mole ratio in the balanced equation.

Solution: First let's calculate the moles of the metal (M) reacted.

$$\text{mol M} = 0.225 \text{ g M} \times \frac{1 \text{ mol M}}{27.0 \text{ g M}} = 8.33 \times 10^{-3} \text{ mol M}$$

Solve the ideal gas equation algebraically for n_{H_2}. Then, calculate the moles of H_2 by substituting the known quantities into the equation.

$$P = 741 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.975 \text{ atm}$$

$$T = 17° + 273° = 290 \text{ K}$$

$$n_{H_2} = \frac{PV_{H_2}}{RT}$$

$$n_{H_2} = \frac{(0.975 \text{ atm})(0.303 \text{ L})}{\left(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(290 \text{ K})} = 1.24 \times 10^{-2} \text{ mol } H_2$$

Compare the number moles of H_2 produced to the number of moles of M reacted.

$$\frac{1.24 \times 10^{-2} \text{ mol } H_2}{8.33 \times 10^{-3} \text{ mol M}} \approx 1.5$$

This means that the mole ratio of H_2 to M is 1.5 : 1.

We can now write the balanced equation since we know the mole ratio between H_2 and M.

The unbalanced equation is:

$$M(s) + HCl(aq) \longrightarrow 1.5H_2(g) + M_xCl_y(aq)$$

We have 3 atoms of H on the products side of the reaction, so a 3 must be placed in front of HCl. The ratio of M to Cl on the reactants side is now 1 : 3. Therefore the formula of the metal chloride must be MCl_3.

The balanced equation is:

$$\mathbf{M(\mathit{s}) + 3HCl(\mathit{aq}) \longrightarrow 1.5H_2(\mathit{g}) + MCl_3(\mathit{aq})}$$

From the formula of the metal chloride, we determine that the charge of the metal is +3. Therefore, the formula of the metal oxide and the metal sulfate are $\mathbf{M_2O_3}$ and $\mathbf{M_2(SO_4)_3}$, respectively.

5.57 The balanced equation for the reaction is: $NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$

First, we must determine which of the two reactants is the limiting reagent. We find the number of moles of each reactant.

$$?\ \text{mol NH}_3 = 73.0\text{ g NH}_3 \times \frac{1\text{ mol NH}_3}{17.03\text{ g NH}_3} = 4.29\text{ mol NH}_3$$

$$?\ \text{mol HCl} = 73.0\text{ g HCl} \times \frac{1\text{ mol HCl}}{36.46\text{ g HCl}} = 2.00\text{ mol HCl}$$

Since NH_3 and HCl react in a 1:1 mole ratio, HCl is the limiting reagent. The mass of NH_4Cl formed is:

$$\mathbf{?\ g\ NH_4Cl} = 2.00\text{ mol HCl} \times \frac{1\text{ mol NH}_4\text{Cl}}{1\text{ mol HCl}} \times \frac{53.49\text{ g NH}_4\text{Cl}}{1\text{ mol NH}_4\text{Cl}} = \mathbf{107\ g\ NH_4Cl}$$

The gas remaining is ammonia, NH_3. The number of moles of NH_3 remaining is (4.29 – 2.00) mol = 2.29 mol NH_3. The volume of NH_3 gas is:

$$\boldsymbol{V}_{\mathbf{NH_3}} = \frac{n_{\text{NH}_3}RT}{P} = \frac{(2.29\text{ mol})\left(0.0821\,\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(14 + 273)\text{K}}{\left(752\text{ mmHg} \times \frac{1\text{ atm}}{760\text{ mmHg}}\right)} = \mathbf{54.5\ L\ NH_3}$$

5.58 From the moles of CO_2 produced, we can calculate the amount of calcium carbonate that must have reacted. We can then determine the percent by mass of $CaCO_3$ in the 3.00 g sample.

The balanced equation is:

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CO_2(g) + CaCl_2(aq) + H_2O(l)$$

The moles of CO_2 produced can be calculated using the ideal gas equation.

$$n_{\text{CO}_2} = \frac{PV_{\text{CO}_2}}{RT}$$

$$n_{\text{CO}_2} = \frac{\left(792\text{ mmHg} \times \frac{1\text{ atm}}{760\text{ mmHg}}\right)(0.656\text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(20 + 273\text{ K})} = \mathbf{2.84 \times 10^{-2}\ mol\ CO_2}$$

The balanced equation shows a 1:1 mole ratio between CO_2 and $CaCO_3$. Therefore, 2.84×10^{-2} mole of $CaCO_3$ must have reacted.

$$? \text{ g CaCO}_3 \text{ reacted} = (2.84\times10^{-2}\text{ mol CaCO}_3)\times\frac{100.1\text{ g CaCO}_3}{1\text{ mol CaCO}_3} = 2.84\text{ g CaCO}_3$$

The percent by mass of the $CaCO_3$ sample is:

$$\textbf{\% CaCO}_\mathbf{3} = \frac{2.84\text{ g}}{3.00\text{ g}}\times100\% = \mathbf{94.7\%}$$

Assumption: The impurity (or impurities) must not react with HCl to produce CO_2 gas.

5.59 The balanced equation is: $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$

At STP, 1 mole of an ideal gas occupies a volume of 22.41 L. We can use this as a conversion factor to find the moles of H_2 reacted. Then, we can calculate the mass of HCl produced.

$$? \text{ mol H}_2\text{ reacted} = 5.6\text{ L H}_2\times\frac{1\text{ mol H}_2}{22.41\text{ L H}_2} = 0.25\text{ mol H}_2$$

The mass of HCl produced is:

$$\textbf{? g HCl} = 0.25\text{ mol H}_2\times\frac{2\text{ mol HCl}}{1\text{ mol H}_2}\times\frac{36.46\text{ g HCl}}{1\text{ mol HCl}} = \mathbf{18\text{ g HCl}}$$

5.60 The balanced equation is:

$$C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$$

The moles of O_2 needed to react with 227 g ethanol are:

$$227\text{ g C}_2\text{H}_5\text{OH}\times\frac{1\text{ mol C}_2\text{H}_5\text{OH}}{46.07\text{ g C}_2\text{H}_5\text{OH}}\times\frac{3\text{ mol O}_2}{1\text{ mol C}_2\text{H}_5\text{OH}} = 14.8\text{ mol O}_2$$

14.8 moles of O_2 correspond to a volume of:

$$V_{O_2} = \frac{n_{O_2}RT}{P} = \frac{(14.8\text{ mol O}_2)\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(35+273\text{ K})}{\left(790\text{ mmHg}\times\frac{1\text{ atm}}{760\text{ mmHg}}\right)} = 3.60\times10^2\text{ L O}_2$$

Since air is 21.0 percent O_2 by volume, we can write:

$$\mathbf{V_{air}} = V_{O_2}\left(\frac{100\%\text{ air}}{21\%\text{ O}_2}\right) = (3.60\times10^2\text{ L O}_2)\left(\frac{100\%\text{ air}}{21\%\text{ O}_2}\right) = \mathbf{1.71\times10^3\text{ L air}}$$

5.63 First, we calculate the mole fraction of each component of the mixture. Then, we can calculate the partial pressure of each component using the equation, $P_i = X_iP_T$.

The number of moles of the combined gases is:

$$n = n_{CH_4} + n_{C_2H_6} + n_{C_3H_8} = 0.31\text{ mol} + 0.25\text{ mol} + 0.29\text{ mol} = 0.85\text{ mol}$$

$$X_{CH_4} = \frac{0.31\text{ mol}}{0.85\text{ mol}} = 0.36 \qquad X_{C_2H_6} = \frac{0.25\text{ mol}}{0.85\text{ mol}} = 0.29 \qquad X_{C_3H_8} = \frac{0.29\text{ mol}}{0.85\text{ mol}} = 0.34$$

The partial pressures are:

$$\boldsymbol{P_{CH_4}} = X_{CH_4} \times P_{total} = 0.36 \times 1.50\text{ atm} = \mathbf{0.54\ atm}$$

$$\boldsymbol{P_{C_2H_6}} = X_{C_2H_6} \times P_{total} = 0.29 \times 1.50\text{ atm} = \mathbf{0.44\ atm}$$

$$\boldsymbol{P_{C_3H_8}} = X_{C_3H_8} \times P_{total} = 0.34 \times 1.50\text{ atm} = \mathbf{0.51\ atm}$$

5.64 Dalton's law states that the total pressure of the mixture is the sum of the partial pressures.

(a) $\boldsymbol{P_{total}} = 0.32\text{ atm} + 0.15\text{ atm} + 0.42\text{ atm} = \mathbf{0.89\ atm}$

(b) We know:

Initial conditions	Final Conditions
$P_1 = (0.15 + 0.42)\text{atm} = 0.57\text{ atm}$	$P_2 = 1.0\text{ atm}$
$T_1 = (15 + 273)\text{K} = 288\text{ K}$	$T_2 = 273\text{ K}$
$V_1 = 2.5\text{ L}$	$V_2 = ?$

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

Because $n_1 = n_2$, we can write:

$$V_2 = \frac{P_1V_1T_2}{P_2T_1}$$

$$\boldsymbol{V_2} = \frac{(0.57\text{ atm})(2.5\text{ L})(273\text{ K})}{(1.0\text{ atm})(288\text{ K})} = \mathbf{1.4\ L\ at\ STP}$$

5.65 Since volume is proportional to the number of moles of gas present, we can directly convert the volume percents to mole fractions.

$$X_{N_2} = 0.7808 \qquad X_{O_2} = 0.2094 \qquad X_{Ar} = 0.0093 \qquad X_{CO_2} = 0.0005$$

(a) For each gas, $P_i = X_iP_T = X_i(1.00\text{ atm})$.

$$\boldsymbol{P_{N_2}} = \mathbf{0.781\ atm}, \quad \boldsymbol{P_{O_2}} = \mathbf{0.209\ atm}, \quad \boldsymbol{P_{Ar}} = \mathbf{9.3 \times 10^{-3}\ atm}, \quad \boldsymbol{P_{CO_2}} = \mathbf{5 \times 10^{-4}\ atm}$$

(b) Concentration (mol/L) is $c = \frac{n}{V} = \frac{P}{RT}$. Therefore, we have:

$$\boldsymbol{c_{N_2}} = \frac{0.781\text{ atm}}{\left(0.0821\,\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273\text{ K})} = \mathbf{3.48 \times 10^{-2}}\ \boldsymbol{M}$$

Similarly, $\boldsymbol{c_{O_2}} = \mathbf{9.32 \times 10^{-3}}\ \boldsymbol{M}$, $\boldsymbol{c_{Ar}} = \mathbf{4.1 \times 10^{-4}}\ \boldsymbol{M}$, $\boldsymbol{c_{CO_2}} = \mathbf{2 \times 10^{-5}}\ \boldsymbol{M}$

5.66 $P_{\text{Total}} = P_1 + P_2 + P_3 + \ldots + P_n$

In this case,

$$P_{\text{Total}} = P_{\text{Ne}} + P_{\text{He}} + P_{\text{H}_2\text{O}}$$

$$P_{\text{Ne}} = P_{\text{Total}} - P_{\text{He}} - P_{\text{H}_2\text{O}}$$

$$\boldsymbol{P_{\textbf{Ne}}} = 745\text{ mm Hg} - 368\text{ mmHg} - 28.3\text{ mmHg} = \mathbf{349\ mmHg}$$

5.67 If we can calculate the moles of H_2 gas collected, we can determine the amount of Na that must have reacted. We can calculate the moles of H_2 gas using the ideal gas equation.

$$P_{\text{H}_2} = P_{\text{Total}} - P_{\text{H}_2\text{O}} = 1.00\text{ atm} - 0.0313\text{ atm} = 0.97\text{ atm}$$

The number of moles of hydrogen gas collected is:

$$n_{\text{H}_2} = \frac{P_{\text{H}_2}V}{RT} = \frac{(0.97\text{ atm})(0.246\text{ L})}{\left(0.0821\,\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(25 + 273)\text{K}} = 0.0098\text{ mol H}_2$$

The balanced equation shows a 2:1 mole ratio between Na and H_2. The mass of Na consumed in the reaction is:

$$\textbf{? g Na} = 0.0098\text{ mol H}_2 \times \frac{2\text{ mol Na}}{1\text{ mol H}_2} \times \frac{22.99\text{ g Na}}{1\text{ mol Na}} = \mathbf{0.45\ g\ Na}$$

5.68 **Strategy:** To solve for moles of H_2 generated, we must first calculate the partial pressure of H_2 in the mixture. What gas law do we need? How do we convert from moles of H_2 to amount of Zn reacted?

Solution: Dalton's law of partial pressure states that

$$P_{\text{Total}} = P_1 + P_2 + P_3 + \ldots + P_n$$

In this case,

$$P_{\text{Total}} = P_{\text{H}_2} + P_{\text{H}_2\text{O}}$$

$$P_{\text{H}_2} = P_{\text{Total}} - P_{\text{H}_2\text{O}}$$

$$P_{\text{H}_2} = 0.980\text{ atm} - (23.8\text{ mmHg})\left(\frac{1\text{ atm}}{760\text{ mmHg}}\right) = 0.949\text{ atm}$$

Now that we know the pressure of H_2 gas, we can calculate the moles of H_2. Then, using the mole ratio from the balanced equation, we can calculate moles of Zn.

$$n_{\text{H}_2} = \frac{P_{\text{H}_2}V}{RT}$$

$$n_{\text{H}_2} = \frac{(0.949\text{ atm})(7.80\text{ L})}{(25 + 273)\text{K}} \times \frac{\text{mol}\cdot\text{K}}{0.0821\text{ L}\cdot\text{atm}} = 0.303\text{ mol H}_2$$

Using the mole ratio from the balanced equation and the molar mass of zinc, we can now calculate the grams of zinc consumed in the reaction.

$$\textbf{? g Zn} = 0.303 \text{ mol } H_2 \times \frac{1 \text{ mol Zn}}{1 \text{ mol } H_2} \times \frac{65.39 \text{ g Zn}}{1 \text{ mol Zn}} = \textbf{19.8 g Zn}$$

5.69 In the mixture, the temperature and volume occupied are the same for the two gases, so the pressure should be proportional to the number of moles. Recall that $P_i = X_iP_T$. The mole fraction of oxygen is:

$$X_{O_2} = \frac{P_{O_2}}{P_{\text{total}}} = \frac{0.20 \text{ atm}}{4.2 \text{ atm}} = 0.048$$

In other words 4.8% of the gas particles are oxygen molecules, which occupy **4.8%** of the volume.

5.70 $P_i = X_iP_T$

We need to determine the mole fractions of each component in order to determine their partial pressures. To calculate mole fraction, write the balanced chemical equation to determine the correct mole ratio.

$$2NH_3(g) \longrightarrow N_2(g) + 3H_2(g)$$

The mole fractions of H_2 and N_2 are:

$$X_{H_2} = \frac{3 \text{ mol}}{3 \text{ mol} + 1 \text{ mol}} = 0.750$$

$$X_{N_2} = \frac{1 \text{ mol}}{3 \text{ mol} + 1 \text{ mol}} = 0.250$$

The partial pressures of H_2 and N_2 are:

$$\boldsymbol{P_{H_2}} = X_{H_2}P_T = (0.750)(866 \text{ mmHg}) = \textbf{650 mmHg}$$

$$\boldsymbol{P_{N_2}} = X_{N_2}P_T = (0.250)(866 \text{ mmHg}) = \textbf{217 mmHg}$$

5.77 $u_{\text{rms}} = \sqrt{\dfrac{3RT}{\mathcal{M}}}$

O_2: $$\boldsymbol{u_{\text{rms}}} = \sqrt{\frac{3(8.314 \text{ J/K}\cdot\text{mol})(65 + 273)\text{K}}{32.00 \times 10^{-3} \text{ kg/mol}}} = \textbf{513 m/s}$$

UF_6: $$\boldsymbol{u_{\text{rms}}} = \sqrt{\frac{3(8.314 \text{ J/K}\cdot\text{mol})(65 + 273)\text{K}}{352.00 \times 10^{-3} \text{ kg/mol}}} = \textbf{155 m/s}$$

As should be the case, the heavier gas, UF_6, has a smaller average velocity than the lighter gas, O_2.

5.78 **Strategy:** To calculate the root-mean-square speed, we use Equation (5.16) of the text. What units should we use for R and $\mathcal{M}$ so the u_{rms} will be expressed in units of m/s.

Solution: To calculate u_{rms}, the units of R should be 8.314 J/mol·K, and because 1 J = 1 $\text{kg}\cdot\text{m}^2/\text{s}^2$, the units of molar mass must be kg/mol.

First, let's calculate the molar masses ($\mathcal{M}$) of N_2, O_2, and O_3. Remember, $\mathcal{M}$ must be in units of kg/mol.

$$\mathcal{M}_{N_2} = 2(14.01\text{ g/mol}) = 28.02\frac{\text{g}}{\text{mol}} \times \frac{1\text{ kg}}{1000\text{ g}} = 0.02802\text{ kg/mol}$$

$$\mathcal{M}_{O_2} = 2(16.00\text{ g/mol}) = 32.00\frac{\text{g}}{\text{mol}} \times \frac{1\text{ kg}}{1000\text{ g}} = 0.03200\text{ kg/mol}$$

$$\mathcal{M}_{O_3} = 3(16.00\text{ g/mol}) = 48.00\frac{\text{g}}{\text{mol}} \times \frac{1\text{ kg}}{1000\text{ g}} = 0.04800\text{ kg/mol}$$

Now, we can substitute into Equation (5.16) of the text.

$$u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

$$u_{\text{rms}}(N_2) = \sqrt{\frac{(3)\left(8.314\frac{\text{J}}{\text{mol}\cdot\text{K}}\right)(-23+273)\text{K}}{\left(0.02802\frac{\text{kg}}{\text{mol}}\right)}}$$

$$\boldsymbol{u_{\text{rms}}(N_2) = 472\text{ m/s}}$$

Similarly,

$$\boldsymbol{u_{\text{rms}}(O_2) = 441\text{ m/s}} \qquad \boldsymbol{u_{\text{rms}}(O_3) = 360\text{ m/s}}$$

Check: Since the molar masses of the gases increase in the order: $N_2 < O_2 < O_3$, we expect the lightest gas (N_2) to move the fastest on average and the heaviest gas (O_3) to move the slowest on average. This is confirmed in the above calculation.

5.79 **(a)** Inversely proportional to density
(b) Independent of temperature
(c) Decreases with increasing pressure
(d) Increases with increasing volume
(e) Inversely proportional to size

5.80 $$\textbf{RMS speed} = \sqrt{\frac{\left(2.0^2+2.2^2+2.6^2+2.7^2+3.3^2+3.5^2\right)(\text{m/s})^2}{6}} = \mathbf{2.8\text{ m/s}}$$

$$\textbf{Average speed} = \frac{(2.0+2.2+2.6+2.7+3.3+3.5)\text{m/s}}{6} = \mathbf{2.7\text{ m/s}}$$

The root-mean-square value is always greater than the average value, because squaring favors the larger values compared to just taking the average value.

5.81 We know that the root-mean-square speed (u_{rms}) of a gas can be calculated as follows:

$$u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

The rate of diffusion (r) will be directly proportional to the root-mean-square speed. Gases moving at greater speeds will diffuse faster. For two different gases we can write the rates of diffusion as follows:

$$r_1 = \sqrt{\frac{3RT}{\mathcal{M}_1}} \qquad r_2 = \sqrt{\frac{3RT}{\mathcal{M}_2}}$$

Dividing r_1 by r_2 gives:

$$\frac{r_1}{r_2} = \frac{\sqrt{\dfrac{3RT}{\mathcal{M}_1}}}{\sqrt{\dfrac{3RT}{\mathcal{M}_2}}}$$

Canceling $3RT$ from the equation gives:

$$\frac{r_1}{r_2} = \sqrt{\frac{\dfrac{1}{\mathcal{M}_1}}{\dfrac{1}{\mathcal{M}_2}}} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$$

Does the derived equation make sense? Assume that gas_1 is a lighter gas (has a smaller molar mass) than gas_2. Dividing a larger molar mass ($\mathcal{M}_2$) by a smaller molar mass ($\mathcal{M}_1$) will give a number larger than 1. This indicates that the lighter gas will diffuse at a faster rate compared to the heavier gas.

5.82 The separation factor is given by:

$$s = \frac{r_1}{r_2} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$$

This equation is the same as Graham's Law, Equation (5.17) of the text. For $^{235}UF_6$ and $^{238}UF_6$, we have:

$$s = \sqrt{\frac{238 + (6)(19.00)}{235 + (6)(19.00)}} = \mathbf{1.0043}$$

This is a very small separation factor, which is why many (thousands) stages of effusion are needed to enrich ^{235}U.

5.83 The rate of effusion is the number of molecules passing through a porous barrier in a given time. The longer it takes, the slower the rate of effusion. Therefore, Equation (5.17) of the text can be written as

$$\frac{r_1}{r_2} = \frac{t_2}{t_1} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$$

where t_1 and t_2 are the times of effusion for gases 1 and 2, respectively.

The molar mass of N_2 is 28.02 g/mol. We write

$$\frac{15.0 \text{ min}}{12.0 \text{ min}} = \sqrt{\frac{\mathcal{M}}{28.02 \text{ g/mol}}}$$

where $\mathcal{M}$ is the molar mass of the unknown gas. Solving for $\mathcal{M}$, we obtain

$$\mathcal{M} = \left(\frac{15.0\text{ min}}{12.0\text{ min}}\right)^2 \times 28.02\text{ g/mol} = \mathbf{43.8\text{ g/mol}}$$

The gas is **carbon dioxide**, **CO_2** (molar mass = 44.01 g/mol). During the fermentation of glucose, ethanol and carbon dioxide are produced.

5.84 The rate of effusion is the number of molecules passing through a porous barrier in a given time. The molar mass of CH_4 is 16.04 g/mol. Using Equation (5.17) of the text, we find the molar mass of $Ni(CO)_x$.

$$\frac{r_1}{r_2} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$$

$$\frac{3.3}{1.0} = \sqrt{\frac{\mathcal{M}_{Ni(CO)_x}}{16.04\text{ g/mol}}}$$

$$10.89 = \frac{\mathcal{M}_{Ni(CO)_x}}{16.04\text{ g/mol}}$$

$$\mathcal{M}_{Ni(CO)_x} = 174.7\text{ g/mol}$$

To find the value of x, we first subtract the molar mass of Ni from 174.7 g/mol.

$$174.7\text{ g} - 58.69\text{ g} = 116.0\text{ g}$$

116.0 g is the mass of CO in 1 mole of the compound. The mass of 1 mole of CO is 28.01 g.

$$\frac{116.0\text{ g}}{28.01\text{ g}} = 4.141 \approx \mathbf{4}$$

This calculation indicates that there are 4 moles of CO in 1 mole of the compound. The value of ***x*** is **4**.

5.89 In this problem, we are comparing the pressure as determined by the van der waals' equation with that determined by the ideal gas equation.

<u>van der waals' equation:</u>

We find the pressure by first solving algebraically for P.

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

where $n = 2.50$ mol, $V = 5.00$ L, $T = 450$ K, $a = 3.59\text{ atm}\cdot\text{L}^2/\text{mol}^2$, and $b = 0.0427$ L/mol

$$\boldsymbol{P} = \frac{(2.50\text{ mol})\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(450\text{ K})}{[(5.00\text{ L}) - (2.50\text{ mol} \times 0.0427\text{ L/mol})]} - \frac{\left(3.59\frac{\text{atm}\cdot\text{L}^2}{\text{mol}^2}\right)(2.50\text{ mol})^2}{(5.00\text{ L})^2} = \mathbf{18.0\text{ atm}}$$

<u>ideal gas equation:</u>

$$\boldsymbol{P} = \frac{nRT}{V} = \frac{(2.50\text{ mol})\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(450\text{ K})}{(5.00\text{ L})} = \mathbf{18.5\ atm}$$

Since the pressure calculated using van der waals' equation is comparable to the pressure calculated using the ideal gas equation, we conclude that CO_2 behaves fairly ideally under these conditions.

5.90 **Strategy:** In this problem we can determine if the gas deviates from ideal behavior, by comparing the ideal pressure with the actual pressure. We can calculate the ideal gas pressure using the ideal gas equation, and then compare it to the actual pressure given in the problem. What temperature unit should we use in the calculation?

Solution: We convert the temperature to units of Kelvin, then substitute the given quantities into the ideal gas equation.

$$T(\text{K}) = 27°\text{C} + 273° = 300\text{ K}$$

$$P = \frac{nRT}{V} = \frac{(10.0\text{ mol})\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(300\text{ K})}{1.50\text{ L}} = 164\text{ atm}$$

Now, we can compare the ideal pressure to the actual pressure by calculating the percent error.

$$\%\text{ error} = \frac{164\text{ atm} - 130\text{ atm}}{130\text{ atm}} \times 100\% = 26.2\%$$

Based on the large percent error, we conclude that under this condition of high pressure, the gas behaves in a **non-ideal** manner.

5.91 **(a)** Neither the amount of gas in the tire nor its volume change appreciably. The pressure is proportional to the temperature. Therefore, as the temperature rises, the pressure increases.

(b) As the paper bag is hit, its volume decreases so that its pressure increases. The popping sound occurs when the bag is broken.

(c) As the balloon rises, the pressure outside decreases steadily, and the balloon expands.

(d) The pressure inside the bulb is greater than 1 atm.

5.92 When *a* and *b* are zero, the van der Waals equation simply becomes the ideal gas equation. In other words, an ideal gas has zero for the *a* and *b* values of the van der Waals equation. It therefore stands to reason that the gas with the smallest values of *a* and *b* will behave most like an ideal gas under a specific set of pressure and temperature conditions. Of the choices given in the problem, the gas with the smallest *a* and *b* values is **Ne** (see Table 5.4).

5.93 You can map out the following strategy to solve for the total volume of gas.

grams nitroglycerin → moles nitroglycerin → moles products → volume of products

$$?\text{ mol products} = 2.6 \times 10^2\text{ g nitroglycerin} \times \frac{1\text{ mol nitroglycerin}}{227.09\text{ g nitroglycerin}} \times \frac{29\text{ mol product}}{4\text{ mol nitroglycerin}} = 8.3\text{ mol}$$

Calculating the volume of products:

$$\mathbf{V_{product}} = \frac{n_{\text{product}}RT}{P} = \frac{(8.3\text{ mol})\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298\text{ K})}{(1.2\text{ atm})} = \mathbf{1.7\times10^2\ L}$$

The relationship between partial pressure and P_{total} is:

$$P_{\text{i}} = X_{\text{i}}P_{\text{T}}$$

Calculate the mole fraction of each gaseous product, then calculate its partial pressure using the equation above.

$$X_{\text{component}} = \frac{\text{moles component}}{\text{total moles all components}}$$

$$X_{\text{CO}_2} = \frac{12\text{ mol CO}_2}{29\text{ mol product}} = 0.41$$

Similarly, $X_{\text{H}_2\text{O}} = 0.34$, $X_{\text{N}_2} = 0.21$, and $X_{\text{O}_2} = 0.034$

$$P_{\text{CO}_2} = X_{\text{CO}_2}P_{\text{T}}$$

$$\mathbf{P_{CO_2}} = (0.41)(1.2\text{ atm}) = \mathbf{0.49\ atm}$$

Similarly, $\mathbf{P_{H_2O}} = \mathbf{0.41\ atm}$, $\mathbf{P_{N_2}} = \mathbf{0.25\ atm}$, and $\mathbf{P_{O_2}} = \mathbf{0.041\ atm}$.

5.94 We need to determine the molar mass of the gas. Comparing the molar mass to the empirical mass will allow us to determine the molecular formula.

$$n = \frac{PV}{RT} = \frac{(0.74\text{ atm})\left(97.2\text{ mL}\times\frac{0.001\text{ L}}{1\text{ mL}}\right)}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(200+273)\text{K}} = 1.85\times10^{-3}\text{ mol}$$

$$\text{molar mass} = \frac{0.145\text{ g}}{1.85\times10^{-3}\text{ mol}} = 78.4\text{ g/mol}$$

The empirical mass of CH = 13.02 g/mol

Since $\frac{78.4\text{ g/mol}}{13.02\text{ g/mol}} = 6.02 \approx 6$, the molecular formula is $(\text{CH})_6$ or $\mathbf{C_6H_6}$.

5.95 **(a)** $\text{NH}_4\text{NO}_2(s) \longrightarrow \text{N}_2(g) + 2\text{H}_2\text{O}(l)$

(b) Map out the following strategy to solve the problem.

volume N_2 → moles N_2 → moles NH_4NO_2 → grams NH_4NO_2

First, calculate the moles of N_2 using the ideal gas equation.

$$T(\text{K}) = 22° + 273° = 295\text{ K}$$

$$V = 86.2\text{ mL}\times\frac{1\text{ L}}{1000\text{ mL}} = 0.0862\text{ L}$$

$$n_{N_2} = \frac{P_{N_2}V}{RT}$$

$$n_{N_2} = \frac{(1.20\text{ atm})(0.0862\text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(295\text{ K})} = 4.27\times10^{-3}\text{ mol}$$

Next, calculate the mass of NH_4NO_2 needed to produce 4.27×10^{-3} mole of N_2.

$$\textbf{? g NH}_4\textbf{NO}_2 = (4.27\times10^{-3}\text{ mol N}_2)\times\frac{1\text{ mol NH}_4\text{NO}_2}{1\text{ mol N}_2}\times\frac{64.05\text{ g NH}_4\text{NO}_2}{1\text{ mol NH}_4\text{NO}_2} = \textbf{0.273 g}$$

5.96 The reaction is: $HCO_3^-(aq) + H^+(aq) \longrightarrow H_2O(l) + CO_2(g)$

The mass of HCO_3^- reacted is:

$$3.29\text{ g tablet}\times\frac{32.5\%\text{ HCO}_3^-}{100\%\text{ tablet}} = 1.07\text{ g HCO}_3^-$$

$$\text{mol CO}_2\text{ produced} = 1.07\text{ g HCO}_3^-\times\frac{1\text{ mol HCO}_3^-}{61.02\text{ g HCO}_3^-}\times\frac{1\text{ mol CO}_2}{1\text{ mol HCO}_3^-} = 0.0175\text{ mol CO}_2$$

$$V_{\mathbf{CO_2}} = \frac{n_{CO_2}RT}{P} = \frac{(0.0175\text{ mol CO}_2)\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(37+273)\text{K}}{(1.00\text{ atm})} = 0.445\text{ L} = \textbf{445 mL}$$

5.97 No, because an ideal gas cannot be liquefied, since the assumption is that there are no intermolecular forces in an ideal gas.

5.98 **(a)** The number of moles of $Ni(CO)_4$ formed is:

$$86.4\text{ g Ni}\times\frac{1\text{ mol Ni}}{58.69\text{ g Ni}}\times\frac{1\text{ mol Ni(CO)}_4}{1\text{ mol Ni}} = 1.47\text{ mol Ni(CO)}_4$$

The pressure of $Ni(CO)_4$ is:

$$P = \frac{nRT}{V} = \frac{(1.47\text{ mol})\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(43+273)\text{K}}{4.00\text{ L}} = \textbf{9.53 atm}$$

(b) $Ni(CO)_4$ decomposes to produce more moles of gas (CO), which increases the pressure.

$$Ni(CO)_4(g) \longrightarrow Ni(s) + 4CO(g)$$

5.99 The partial pressure of carbon dioxide is higher in the winter because carbon dioxide is utilized less by photosynthesis in plants.

5.100 Using the ideal gas equation, we can calculate the moles of gas.

$$n = \frac{PV}{RT} = \frac{(1.1\text{ atm})\left(5.0\times 10^{2}\text{ mL}\times\dfrac{0.001\text{ L}}{1\text{ mL}}\right)}{\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(37+273)\text{K}} = 0.0216\text{ mol gas}$$

Next, use Avogadro's number to convert to molecules of gas.

$$0.0216\text{ mol gas}\times\frac{6.022\times 10^{23}\text{ molecules}}{1\text{ mol gas}} = \mathbf{1.30\times 10^{22}\text{ molecules of gas}}$$

The most common gases present in exhaled air are: $\mathbf{CO_2}$, $\mathbf{O_2}$, $\mathbf{N_2}$, and $\mathbf{H_2O}$.

5.101 **(a)** Write a balanced chemical equation.

$$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

First, calculate the moles of CO_2 produced.

$$?\text{ mol CO}_2 = 5.0\text{ g NaHCO}_3\times\frac{1\text{ mol NaHCO}_3}{84.01\text{ g NaHCO}_3}\times\frac{1\text{ mol CO}_2}{2\text{ mol NaHCO}_3} = 0.030\text{ mol}$$

Next, calculate the volume of CO_2 produced using the ideal gas equation.

$T(\text{K}) = 180° + 273° = 453\text{ K}$

$$V_{CO_2} = \frac{n_{CO_2}RT}{P}$$

$$\mathbf{V_{CO_2}} = \frac{(0.030\text{ mol})\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(453\text{ K})}{(1.3\text{ atm})} = \mathbf{0.86\text{ L}}$$

(b) The balanced chemical equation for the decomposition of NH_4HCO_3 is

$$NH_4HCO_3(s) \longrightarrow NH_3(g) + CO_2(g) + H_2O(g)$$

The advantage in using the ammonium salt is that more gas is produced per gram of reactant. The disadvantage is that one of the gases is ammonia. The strong odor of ammonia would ***not*** make the ammonium salt a good choice for baking.

5.102 Mass of the Earth's atmosphere = (surface area of the earth in cm^2) × (mass per 1 cm^2 column)

Mass of a single column of air with a surface area of 1 cm^2 area is:

$$76.0\text{ cm}\times 13.6\text{ g/cm}^3 = 1.03\times 10^{3}\text{ g/cm}^2$$

The surface area of the Earth in cm^2 is:

$$4\pi r^2 = 4\pi(6.371\times 10^{8}\text{ cm})^2 = 5.10\times 10^{18}\text{ cm}^2$$

$$\textbf{Mass of atmosphere} = (5.10\times 10^{18}\text{ cm}^2)(1.03\times 10^{3}\text{ g/cm}^2) = 5.25\times 10^{21}\text{ g} = \mathbf{5.25\times 10^{18}\text{ kg}}$$

5.103 First, calculate the moles of H_2 formed.

$$? \text{ mol } H_2 = 3.12 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g}} \times \frac{3 \text{ mol } H_2}{2 \text{ mol Al}} = 0.173 \text{ mol}$$

Next, calculate the volume of H_2 produced using the ideal gas equation.

$$\boldsymbol{V_{H_2}} = \frac{n_{H_2}RT}{P} = \frac{(0.173 \text{ mol})\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(296 \text{ K})}{(1.00 \text{ atm})} = \mathbf{4.20\ L}$$

5.104 To calculate the molarity of NaOH, we need moles of NaOH and volume of the NaOH solution. The volume is given in the problem; therefore, we need to calculate the moles of NaOH. The moles of NaOH can be calculated from the reaction of NaOH with HCl. The balanced equation is:

$$\text{NaOH}(aq) + \text{HCl}(aq) \longrightarrow \text{H}_2\text{O}(l) + \text{NaCl}(aq)$$

The number of moles of HCl gas is found from the ideal gas equation. $V = 0.189$ L, $T = (25 + 273)\text{K} = 298$ K, and $P = 108 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.142 \text{ atm}$.

$$n_{HCl} = \frac{PV_{HCl}}{RT} = \frac{(0.142 \text{ atm})(0.189 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298 \text{ K})} = 1.10 \times 10^{-3} \text{ mol HCl}$$

The moles of NaOH can be calculated using the mole ratio from the balanced equation.

$$(1.10 \times 10^{-3} \text{ mol HCl}) \times \frac{1 \text{ mol NaOH}}{1 \text{ mol HCl}} = 1.10 \times 10^{-3} \text{ mol NaOH}$$

The molarity of the NaOH solution is:

$$\boldsymbol{M} = \frac{\text{mol NaOH}}{\text{L of soln}} = \frac{1.10 \times 10^{-3} \text{ mol NaOH}}{0.0157 \text{ L soln}} = 0.0701 \text{ mol/L} = \mathbf{0.0701}\ \boldsymbol{M}$$

5.105 **(a)** $C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$

(b) From the balanced equation, we see that there is a 1:3 mole ratio between C_3H_8 and CO_2.

$$\mathbf{?\ L\ CO_2} = 7.45 \text{ g } C_3H_8 \times \frac{1 \text{ mol } C_3H_8}{44.09 \text{ g } C_3H_8} \times \frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8} \times \frac{22.414 \text{ L } CO_2}{1 \text{ mol } CO_2} = \mathbf{11.4\ L\ CO_2}$$

5.106 To calculate the partial pressures of He and Ne, the total pressure of the mixture is needed. To calculate the total pressure of the mixture, we need the total number of moles of gas in the mixture (mol He + mol Ne).

$$n_{He} = \frac{PV}{RT} = \frac{(0.63 \text{ atm})(1.2 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(16 + 273)\text{K}} = 0.032 \text{ mol He}$$

$$n_{Ne} = \frac{PV}{RT} = \frac{(2.8 \text{ atm})(3.4 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(16 + 273)\text{K}} = 0.40 \text{ mol Ne}$$

The total pressure is:

$$P_{\text{Total}} = \frac{(n_{\text{He}} + n_{\text{Ne}})RT}{V_{\text{Total}}} = \frac{(0.032 + 0.40)\text{mol}\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(16 + 273)\text{K}}{(1.2 + 3.4)\text{L}} = 2.2 \text{ atm}$$

$P_{\text{i}} = X_{\text{i}}P_{\text{T}}$. The partial pressures of He and Ne are:

$$\boldsymbol{P_{\text{He}}} = \frac{0.032 \text{ mol}}{(0.032 + 0.40)\text{mol}} \times 2.2 \text{ atm} = \mathbf{0.16\ atm}$$

$$\boldsymbol{P_{\text{Ne}}} = \frac{0.40 \text{ mol}}{(0.032 + 0.40)\text{mol}} \times 2.2 \text{ atm} = \mathbf{2.0\ atm}$$

5.107 Calculate the initial number of moles of NO and O_2 using the ideal gas equation.

$$n_{\text{NO}} = \frac{P_{\text{NO}}V}{RT} = \frac{(0.500 \text{ atm})(4.00 \text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298 \text{ K})} = 0.0817 \text{ mol NO}$$

$$n_{O_2} = \frac{P_{O_2}V}{RT} = \frac{(1.00 \text{ atm})(2.00 \text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298 \text{ K})} = 0.0817 \text{ mol } O_2$$

Determine which reactant is the limiting reagent. The number of moles of NO and O_2 calculated above are equal; however, the balanced equation shows that twice as many moles of NO are needed compared to O_2. Thus, NO is the limiting reagent.

Determine the molar amounts of NO, O_2, and NO_2 after complete reaction.

mol NO = 0 mol (All NO is consumed during reaction)

$$\textbf{mol NO}_\mathbf{2} = 0.0817 \text{ mol NO} \times \frac{2 \text{ mol } NO_2}{2 \text{ mol NO}} = \mathbf{0.0817\ mol\ NO_2}$$

$$\text{mol } O_2 \text{ consumed} = 0.0817 \text{ mol NO} \times \frac{1 \text{ mol } O_2}{2 \text{ mol NO}} = 0.0409 \text{ mol } O_2 \text{ consumed}$$

mol O_2 remaining = 0.0817 mol O_2 initial – 0.0409 mol O_2 consumed = **0.0408 mol O_2**

Calculate the partial pressures of O_2 and NO_2 using the ideal gas equation.

Volume of entire apparatus = 2.00 L + 4.00 L = 6.00 L
T(K) = 25° + 273° = 298 K

$$\boldsymbol{P_{O_2}} = \frac{n_{O_2}RT}{V} = \frac{(0.0408 \text{ mol})\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298 \text{ K})}{(6.00 \text{ L})} = \mathbf{0.166\ atm}$$

$$\boldsymbol{P_{NO_2}} = \frac{n_{NO_2}RT}{V} = \frac{(0.0817 \text{ mol})\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298 \text{ K})}{(6.00 \text{ L})} = \mathbf{0.333\ atm}$$

5.108 When the water enters the flask from the dropper, some hydrogen chloride dissolves, creating a partial vacuum. Pressure from the atmosphere forces more water up the vertical tube.

5.109 **(a)** First, the total pressure (P_{Total}) of the mixture of carbon dioxide and hydrogen must be determined at a given temperature in a container of known volume. Next, the carbon dioxide can be removed by reaction with sodium hydroxide.

$$CO_2(g) + 2NaOH(aq) \longrightarrow Na_2CO_3(aq) + H_2O(l)$$

The pressure of the hydrogen gas that remains can now be measured under the same conditions of temperature and volume. Finally, the partial pressure of CO_2 can be calculated.

$$P_{CO_2} = P_{\text{Total}} - P_{H_2}$$

(b) The most direct way to measure the partial pressures would be to use a mass spectrometer to measure the mole fractions of the gases. The partial pressures could then be calculated from the mole fractions and the total pressure. Another way to measure the partial pressures would be to realize that helium has a much lower boiling point than nitrogen. Therefore, nitrogen gas can be removed by lowering the temperature until nitrogen liquefies. Helium will remain as a gas. As in part (a), the total pressure is measured first. Then, the pressure of helium can be measured after the nitrogen is removed. Finally, the pressure of nitrogen is simply the difference between the total pressure and the pressure of helium.

5.110 Use the ideal gas equation to calculate the moles of water produced. We carry an extra significant figure in the first step of the calculation to limit rounding errors.

$$n_{H_2O} = \frac{PV}{RT} = \frac{(24.8\text{ atm})(2.00\text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(120 + 273)\text{K}} = 1.537\text{ mol } H_2O$$

Next, we can determine the mass of H_2O in the 54.2 g sample. Subtracting the mass of H_2O from 54.2 g will give the mass of $MgSO_4$ in the sample.

$$1.537\text{ mol } H_2O \times \frac{18.02\text{ g } H_2O}{1\text{ mol } H_2O} = 27.7\text{ g } H_2O$$

$$\text{Mass } MgSO_4 = 54.2\text{ g sample} - 27.7\text{ g } H_2O = 26.5\text{ g } MgSO_4$$

Finally, we can calculate the moles of $MgSO_4$ in the sample. Comparing moles of $MgSO_4$ to moles of H_2O will allow us to determine the correct mole ratio in the formula.

$$26.5\text{ g } MgSO_4 \times \frac{1\text{ mol } MgSO_4}{120.4\text{ g } MgSO_4} = 0.220\text{ mol } MgSO_4$$

$$\frac{\text{mol } H_2O}{\text{mol } MgSO_4} = \frac{1.54\text{ mol}}{0.220\text{ mol}} = 7.00$$

Therefore, the mole ratio between H_2O and $MgSO_4$ in the compound is 7 : 1. Thus, the value of ***x* = 7**, and the formula is **$MgSO_4\cdot 7H_2O$**.

5.111 The reactions are:

$$Na_2CO_3(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

$$MgCO_3(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2O(l) + CO_2(g)$$

First, let's calculate the moles of CO_2 produced using the ideal gas equation. We carry an extra significant figure throughout this calculation to limit rounding errors.

$$n_{CO_2} = \frac{PV}{RT} = \frac{(1.24\text{ atm})(1.67\text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(26 + 273)\text{K}} = 0.08436\text{ mol CO}_2$$

Since there is a 1:1 mole ratio between CO_2 and both reactants (Na_2CO_3 and $MgCO_3$), then 0.08436 mole of the mixture must have reacted.

$$\text{mol Na}_2\text{CO}_3 + \text{mol MgCO}_3 = 0.08436\text{ mol}$$

Let x be the mass of Na_2CO_3 in the mixture, then $(7.63 - x)$ is the mass of $MgCO_3$ in the mixture.

$$\left[x\text{ g Na}_2\text{CO}_3 \times \frac{1\text{ mol Na}_2\text{CO}_3}{106.0\text{ g Na}_2\text{CO}_3}\right] + \left[(7.63 - x)\text{g MgCO}_3 \times \frac{1\text{ mol MgCO}_3}{84.32\text{ g MgCO}_3}\right] = 0.08436\text{ mol}$$

$$0.009434x - 0.01186x + 0.09049 = 0.08436$$

$$x = 2.527\text{ g} = \text{mass of Na}_2\text{CO}_3\text{ in the mixture}$$

The percent composition by mass of Na_2CO_3 in the mixture is:

$$\textbf{mass \% Na}_2\textbf{CO}_3 = \frac{\text{mass Na}_2\text{CO}_3}{\text{mass of mixture}} \times 100\% = \frac{2.527\text{ g}}{7.63\text{ g}} \times 100\% = \mathbf{33.1\%\ Na_2CO_3}$$

5.112 The circumference of the cylinder is $= 2\pi r = 2\pi\left(\frac{15.0\text{ cm}}{2}\right) = 47.1\text{ cm}$

(a) The speed at which the target is moving equals:

$$\text{speed of target} = \text{circumference} \times \text{revolutions/sec}$$

$$\textbf{speed of target} = \frac{47.1\text{ cm}}{1\text{ revolution}} \times \frac{130\text{ revolutions}}{1\text{ s}} \times \frac{0.01\text{ m}}{1\text{ cm}} = \mathbf{61.2\ m/s}$$

(b) $2.80\text{ cm} \times \frac{0.01\text{ m}}{1\text{ cm}} \times \frac{1\text{ s}}{61.2\text{ m}} = \mathbf{4.58 \times 10^{-4}\ s}$

(c) The Bi atoms must travel across the cylinder to hit the target. This distance is the diameter of the cylinder, which is 15.0 cm. The Bi atoms travel this distance in 4.58×10^{-4} s.

$$\frac{15.0\text{ cm}}{4.58 \times 10^{-4}\text{ s}} \times \frac{0.01\text{ m}}{1\text{ cm}} = \mathbf{328\ m/s}$$

$$u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}} = \sqrt{\frac{3(8.314\text{ J/K}\cdot\text{mol})(850 + 273)\text{K}}{209.0 \times 10^{-3}\text{ kg/mol}}} = \mathbf{366\ m/s}$$

The magnitudes of the speeds are comparable, but not identical. This is not surprising since 328 m/s is the velocity of a particular Bi atom, and u_{rms} is an average value.

5.113 Using the ideal gas equation, we can calculate the moles of water that would be vaporized. We can then convert to mass of water vaporized.

$$n = \frac{PV}{RT} = \frac{\left(187.5\text{ mmHg} \times \dfrac{1\text{ atm}}{760\text{ mmHg}}\right)(2.500\text{ L})}{\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(65 + 273)\text{K}} = 0.0222\text{ mol H}_2\text{O}$$

$$\textbf{? g H}_\mathbf{2}\textbf{O vaporized} = 0.0222\text{ mol H}_2\text{O} \times \frac{18.02\text{ g H}_2\text{O}}{1\text{ mol H}_2\text{O}} = \mathbf{0.400\text{ g H}_2\text{O}}$$

5.114 The moles of O_2 can be calculated from the ideal gas equation. The mass of O_2 can then be calculated using the molar mass as a conversion factor.

$$n_{\text{O}_2} = \frac{PV}{RT} = \frac{(132\text{ atm})(120\text{ L})}{\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(22 + 273)\text{K}} = 654\text{ mol O}_2$$

$$\textbf{? g O}_\mathbf{2} = 654\text{ mol O}_2 \times \frac{32.00\text{ g O}_2}{1\text{ mol O}_2} = \mathbf{2.09 \times 10^4\text{ g O}_2}$$

The volume of O_2 gas under conditions of 1.00 atm pressure and a temperature of 22°C can be calculated using the ideal gas equation. The moles of O_2 = 654 moles.

$$V_{\text{O}_2} = \frac{n_{\text{O}_2}RT}{P} = \frac{(654\text{ mol})\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(22 + 273)\text{K}}{1.00\text{ atm}} = \mathbf{1.58 \times 10^4\text{ L O}_2}$$

5.115 The air inside the egg expands with increasing temperature. The increased pressure can cause the egg to crack.

5.116 The fruit ripens more rapidly because the quantity (partial pressure) of ethylene gas inside the bag increases.

5.117 The balanced equation is: $CO_2(g) + 2NH_3(g) \longrightarrow (NH_2)_2CO(s) + H_2O(g)$

First, we can calculate the moles of NH_3 needed to produce 1.0 ton of urea. Then, we can use the ideal gas equation to calculate the volume of NH_3.

$$\text{? mol NH}_3 = 1.0\text{ ton urea} \times \frac{2000\text{ lb}}{1\text{ ton}} \times \frac{453.6\text{ g}}{1\text{ lb}} \times \frac{1\text{ mol urea}}{60.06\text{ g urea}} \times \frac{2\text{ mol NH}_3}{1\text{ mol urea}} = 3.0 \times 10^4\text{ mol NH}_3$$

$$V_{\text{NH}_3} = \frac{nRT}{P} = \frac{(3.0 \times 10^4\text{ mol})\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(200 + 273)\text{K}}{150\text{ atm}} = \mathbf{7.8 \times 10^3\text{ L NH}_3}$$

5.118 As the pen is used the amount of ink decreases, increasing the volume inside the pen. As the volume increases, the pressure inside the pen decreases. The hole is needed to equalize the pressure as the volume inside the pen increases.

5.119 **(a)** This is a Boyle's Law problem, pressure and volume vary. Assume that the pressure at the water surface is 1 atm. The pressure that the diver experiences 36 ft below water is:

$$1\text{ atm} + \left(36\text{ ft} \times \frac{1\text{ atm}}{33\text{ ft}}\right) = 2.1\text{ atm}$$

$$P_1V_1 = P_2V_2 \qquad \text{or} \qquad \frac{V_1}{V_2} = \frac{P_2}{P_1}$$

$$\frac{V_1}{V_2} = \frac{2.1\text{ atm}}{1\text{ atm}} = \mathbf{2.1}$$

The diver's lungs would increase in volume **2.1 times** by the time he reaches the surface.

(b) $P_{O_2} = X_{O_2}P_T$

$$P_{O_2} = \frac{n_{O_2}}{n_{O_2} + n_{N_2}}P_T$$

At constant temperature and pressure, the volume of a gas is directly proportional to the number of moles of gas. We can write:

$$P_{O_2} = \frac{V_{O_2}}{V_{O_2} + V_{N_2}}P_T$$

We know the partial pressure of O_2 in air, and we know the total pressure exerted on the diver. Plugging these values into the above equation gives:

$$0.20\text{ atm} = \frac{V_{O_2}}{V_{O_2} + V_{N_2}}(4.00\text{ atm})$$

$$\frac{V_{O_2}}{V_{O_2} + V_{N_2}} = \frac{0.20\text{ atm}}{4.00\text{ atm}} = \mathbf{0.050}$$

In other words, the air that the diver breathes must have an oxygen content of **5% by volume**.

5.120 **(a)** $NH_4NO_3(s) \longrightarrow N_2O(g) + 2H_2O(l)$

(b) $$\boldsymbol{R} = \frac{PV}{nT} = \frac{\left(718\text{ mmHg} \times \frac{1\text{ atm}}{760\text{ mmHg}}\right)(0.340\text{ L})}{\left(0.580\text{ g N}_2\text{O} \times \frac{1\text{ mol N}_2\text{O}}{44.02\text{ g N}_2\text{O}}\right)(24 + 273)\text{K}} = \mathbf{0.0821\frac{L\cdot atm}{mol\cdot K}}$$

5.121 Since Ne and NH_3 are at the same temperature, they have the same average kinetic energy.

$$\text{KE} = \frac{1}{2}m\overline{u^2}$$

or

$$\overline{u^2} = \frac{2\text{KE}}{m}$$

Recall that mass must have units of kg because kinetic energy has units of Joules. $\left(1\text{ J} = \dfrac{1\text{ kg}\cdot\text{m}^2}{\text{s}^2}\right)$

We need to calculate the mass of one Ne atom in kg.

$$\frac{20.18\text{ g Ne}}{1\text{ mol Ne}} \times \frac{1\text{ mol Ne}}{6.022\times10^{23}\text{ Ne atoms}} \times \frac{1\text{ kg}}{1000\text{ g}} = 3.351\times10^{-26}\text{ kg/Ne atom}$$

Solving for $\overline{u^2}$:

$$\overline{\boldsymbol{u}^{\mathbf{2}}} = \frac{2\text{KE}}{m} = \frac{2(7.1\times10^{-21}\text{ J/atom})}{3.351\times10^{-26}\text{ kg/atom}} = \mathbf{4.2\times10^{5}\ m^2/s^2}$$

5.122 The value of *a* indicates how strongly molecules of a given type of gas attract one anther. $\mathbf{C_6H_6}$ has the greatest intermolecular attractions due to its larger size compared to the other choices. Therefore, it has the largest *a* value.

5.123 Using the ideal gas equation, we can calculate the moles of gas in the syringe. Then, knowing the mass of the gas, we can calculate the molar mass of the gas. Finally, comparing the molar mass to the empirical mass will allow us to determine the molecular formula of the gas.

$$n = \frac{PV}{RT} = \frac{(1\text{ atm})(5.58\times10^{-3}\text{ L})}{\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(45+273)\text{K}} = 2.14\times10^{-4}\text{ mol}$$

$$\textbf{molar mass} = \frac{\text{mass (in g)}}{\text{mol}} = \frac{0.0184\text{ g}}{2.14\times10^{-4}\text{ mol}} = \mathbf{86.0\ g/mol}$$

The empirical molar mass of CH_2 is 14.0 g/mol. Dividing the molar mass by the empirical molar mass gives:

$$\frac{86.0\text{ g/mol}}{14.0\text{ g/mol}} \approx 6$$

Therefore, the molecular formula is $(CH_2)_6$ or $\mathbf{C_6H_{12}}$.

5.124 The gases inside the mine were a mixture of carbon dioxide, carbon monoxide, methane, and other harmful compounds. The low atmospheric pressure caused the gases to flow out of the mine (the gases in the mine were at a higher pressure), and the man suffocated.

5.125 **(a)** $CaO(s) + CO_2(g) \longrightarrow CaCO_3(s)$

$BaO(s) + CO_2(g) \longrightarrow BaCO_3(s)$

(b) First, we need to find the number of moles of CO_2 consumed in the reaction. We can do this by calculating the initial moles of CO_2 in the flask and then comparing it to the CO_2 remaining after the reaction.

$$\underline{\text{Initially}}: \quad n_{CO_2} = \frac{PV}{RT} = \frac{\left(746\text{ mmHg} \times \frac{1\text{ atm}}{760\text{ mmHg}}\right)(1.46\text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(35 + 273)\text{K}} = 0.0567\text{ mol }CO_2$$

$$\underline{\text{Remaining}}: \quad n_{CO_2} = \frac{PV}{RT} = \frac{\left(252\text{ mmHg} \times \frac{1\text{ atm}}{760\text{ mmHg}}\right)(1.46\text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(35 + 273)\text{K}} = 0.0191\text{ mol }CO_2$$

Thus, the amount of CO_2 consumed in the reaction is: (0.0567 mol − 0.0191 mol) = 0.0376 mol CO_2.

Since the mole ratio between CO_2 and both reactants (CaO and BaO) is 1:1, 0.0376 mole of the mixture must have reacted. We can write:

mol CaO + mol BaO = 0.0376 mol

Let x = mass of CaO in the mixture, then $(4.88 - x)$ = mass of BaO in the mixture. We can write:

$$\left[x\text{ g CaO} \times \frac{1\text{ mol CaO}}{56.08\text{ g CaO}}\right] + \left[(4.88 - x)\text{g BaO} \times \frac{1\text{ mol BaO}}{153.3\text{ g BaO}}\right] = 0.0376\text{ mol}$$

$$0.01783x - 0.006523x + 0.0318 = 0.0376$$

x = 0.513 g = mass of CaO in the mixture

mass of BaO in the mixture = $4.88 - x$ = 4.37 g

The percent compositions by mass in the mixture are:

CaO: $\frac{0.513\text{ g}}{4.88\text{ g}} \times 100\% = \mathbf{10.5\%}$ **BaO**: $\frac{4.37\text{ g}}{4.88\text{ g}} \times 100\% = \mathbf{89.5\%}$

5.126 **(a)** This is a Boyle's law problem.

$$P_{\text{tire}}V_{\text{tire}} = P_{\text{air}}V_{\text{air}}$$

$$(5.0\text{ atm})(0.98\text{ L}) = (1.0\text{ atm})V_{\text{air}}$$

$$\mathbf{V_{air} = 4.90\ L}$$

(b) Pressure in the tire − atmospheric pressure = gauge pressure

Pressure in the tire − 1.0 atm = 5.0 atm

Pressure in the tire = 6.0 atm

(c) Again, this is a Boyle's law problem.

$$P_{\text{pump}}V_{\text{pump}} = P_{\text{gauge}}V_{\text{gauge}}$$

$$(1\text{ atm})(0.33V_{\text{tire}}) = P_{\text{gauge}}V_{\text{gauge}}$$

$$P_{\text{gauge}} = 0.33\text{ atm}$$

This is the gauge pressure after one pump stroke. After three strokes, the gauge pressure will be (3 × 0.33 atm), or approximately **1 atm**. This is assuming that the initial gauge pressure was zero.

5.127 **(a)** $$\frac{188\text{ g CO}}{1\text{ hr}} \times \frac{1\text{ mol CO}}{28.01\text{ g CO}} \times \frac{1\text{ hr}}{60\text{ min}} = \mathbf{0.112\ mol\ CO/min}$$

(b) 1000 ppm means that there are 1000 particles of gas per 1,000,000 particles of air. The pressure of a gas is directly proportional to the number of particles of gas. We can calculate the partial pressure of CO in atmospheres, assuming that atmospheric pressure is 1 atm.

$$\frac{1000\text{ particles}}{1{,}000{,}000\text{ particles}} \times 1\text{ atm} = 1.0 \times 10^{-3}\text{ atm}$$

A partial pressure of 1.0×10^{-3} atm CO is lethal.

The volume of the garage (in L) is:

$$(6.0\text{ m} \times 4.0\text{ m} \times 2.2\text{ m}) \times \left(\frac{1\text{ cm}}{0.01\text{ m}}\right)^3 \times \frac{1\text{ L}}{1000\text{ cm}^3} = 5.3 \times 10^4\text{ L}$$

From part (a), we know the rate of CO production per minute. In one minute the partial pressure of CO will be:

$$P_{\text{CO}} = \frac{nRT}{V} = \frac{(0.112\text{ mol})\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(20 + 273)\text{K}}{5.3 \times 10^4\text{ L}} = 5.1 \times 10^{-5}\text{ atm CO/min}$$

How many minutes will it take for the partial pressure of CO to reach the lethal level, 1.0×10^{-3} atm?

$$\textbf{? min} = (1.0 \times 10^{-3}\text{ atm CO}) \times \frac{1\text{ min}}{5.1 \times 10^{-5}\text{ atm CO}} = \mathbf{2.0 \times 10^1\ min}$$

5.128 **(a)** First, let's convert the concentration of hydrogen from atoms/cm^3 to mol/L. The concentration in mol/L can be substituted into the ideal gas equation to calculate the pressure of hydrogen.

$$\frac{1\text{ H atom}}{1\text{ cm}^3} \times \frac{1\text{ mol H}}{6.022 \times 10^{23}\text{ H atoms}} \times \frac{1\text{ cm}^3}{1\text{ mL}} \times \frac{1\text{ mL}}{0.001\text{ L}} = \frac{2 \times 10^{-21}\text{ mol H}}{\text{L}}$$

The pressure of H is:

$$P = \left(\frac{n}{V}\right)RT = \left(\frac{2 \times 10^{-21}\text{ mol}}{1\text{ L}}\right)\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(3\text{ K}) = \mathbf{5 \times 10^{-22}\ atm}$$

(b) From part (a), we know that 1 L contains 1.66×10^{-21} mole of H atoms. We convert to the volume that contains 1.0 g of H atoms.

$$\frac{1\text{ L}}{2 \times 10^{-21}\text{ mol H}} \times \frac{1\text{ mol H}}{1.008\text{ g H}} = \mathbf{5 \times 10^{20}\ L/g\ of\ H}$$

Note: This volume is about that of all the water on Earth!

5.129 **(a)** First, convert density to units of g/L.

$$\frac{0.426\text{ kg}}{1\text{ m}^3} \times \frac{1000\text{ g}}{1\text{ kg}} \times \left(\frac{0.01\text{ m}}{1\text{ cm}}\right)^3 \times \frac{1000\text{ cm}^3}{1\text{ L}} = 0.426\text{ g/L}$$

Let's assume a volume of 1.00 L of air. This air sample will have a mass of 0.426 g. Converting to moles of air:

$$0.426\text{ g air} \times \frac{1\text{ mol air}}{29.0\text{ g air}} = 0.0147\text{ mol air}$$

Now, we can substitute into the ideal gas equation to calculate the air temperature.

$$\boldsymbol{T} = \frac{PV}{nR} = \frac{\left(210\text{ mmHg} \times \dfrac{1\text{ atm}}{760\text{ mmHg}}\right)(1.00\text{ L})}{(0.0147\text{ mol})\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)} = \mathbf{229\text{ K}} = \mathbf{-44°C}$$

(b) To determine the percent decrease in oxygen gas, let's compare moles of O_2 at the top of Mt. Everest to the moles of O_2 at sea level.

$$\frac{n_{O_2}(\text{Mt. Everest})}{n_{O_2}(\text{sea level})} = \frac{\dfrac{P_{O_2}(\text{Mt. Everest})V}{RT}}{\dfrac{P_{O_2}(\text{sea level})V}{RT}}$$

$$\frac{n_{O_2}(\text{Mt. Everest})}{n_{O_2}(\text{sea level})} = \frac{P_{O_2}(\text{Mt. Everest})}{P_{O_2}(\text{sea level})} = \frac{210\text{ mmHg}}{760\text{ mmHg}} = 0.276$$

This calculation indicates that there is only 27.6% as much oxygen at the top of Mt. Everest compared to sea level. Therefore, the percent decrease in oxygen gas from sea level to the top of Mt. Everest is **72.4%**.

5.130 From Table 5.3, the equilibrium vapor pressure at 30°C is 31.82 mmHg.

Converting 3.9×10^3 Pa to units of mmHg:

$$(3.9 \times 10^3\text{ Pa}) \times \frac{760\text{ mmHg}}{1.01325 \times 10^5\text{ Pa}} = 29\text{ mmHg}$$

$$\textbf{Relative Humidity} = \frac{\text{partial pressure of water vapor}}{\text{equilibrium vapor pressure}} \times 100\% = \frac{29\text{ mmHg}}{31.82\text{ mmHg}} \times 100\% = \mathbf{91\%}$$

5.131 At the same temperature and pressure, the same volume contains the same number of moles of gases. Since water has a lower molar mass (18.02 g/mol) than air (about 29 g/mol), moisture laden air weighs less than dry air.

Under conditions of constant temperature and volume, moist air exerts a lower pressure than dry air. Hence, a low-pressure front indicates that the air is moist.

5.132 The volume of one alveoli is:

$$V = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi(0.0050\text{ cm})^3 = (5.2\times10^{-7}\text{ cm}^3)\times\frac{1\text{ mL}}{1\text{ cm}^3}\times\frac{0.001\text{ L}}{1\text{ mL}} = 5.2\times10^{-10}\text{ L}$$

The number of moles of air in one alveoli can be calculated using the ideal gas equation.

$$n = \frac{PV}{RT} = \frac{(1.0\text{ atm})(5.2\times10^{-10}\text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(37+273)\text{K}} = 2.0\times10^{-11}\text{ mol of air}$$

Since the air inside the alveoli is 14 percent oxygen, the moles of oxygen in one alveoli equals:

$$(2.0\times10^{-11}\text{ mol of air})\times\frac{14\%\text{ oxygen}}{100\%\text{ air}} = 2.8\times10^{-12}\text{ mol O}_2$$

Converting to O_2 molecules:

$$(2.8\times10^{-12}\text{ mol O}_2)\times\frac{6.022\times10^{23}\text{ O}_2\text{ molecules}}{1\text{ mol O}_2} = \mathbf{1.7\times10^{12}\ O_2\ molecules}$$

5.133 **(a)** We can calculate the moles of mercury vapor using the ideal gas equation, but first we need to know the volume of the room in liters.

$$V_{\text{room}} = (15.2\text{ m})(6.6\text{ m})(2.4\text{ m})\times\left(\frac{1\text{ cm}}{0.01\text{ m}}\right)^3\times\frac{1\text{ L}}{1000\text{ cm}^3} = 2.4\times10^5\text{ L}$$

$$n_{\text{Hg}} = \frac{PV}{RT} = \frac{(1.7\times10^{-6}\text{ atm})(2.4\times10^5\text{ L})}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(20+273)\text{K}} = 0.017\text{ mol Hg}$$

Converting to mass:

$$\mathbf{?\ g\ Hg} = 0.017\text{ mol Hg}\times\frac{200.6\text{ g Hg}}{1\text{ mol Hg}} = \mathbf{3.4\ g\ Hg}$$

(b) The concentration of Hg vapor in the room is:

$$\frac{3.4\text{ g Hg}}{(15.2\text{ m})(6.6\text{ m})(2.4\text{ m})} = 0.014\text{ g Hg/m}^3 = \mathbf{14\ mg\ Hg/m^3}$$

Yes, this far exceeds the air quality regulation of 0.050 mg Hg/m^3 of air.

(c) Physical: The sulfur powder covers the Hg surface, thus retarding the rate of evaporation.

Chemical: Sulfur reacts slowly with Hg to form HgS. HgS has no measurable vapor pressure.

5.134 The molar mass of a gas can be calculated using Equation (5.12) of the text.

$$\mathcal{M} = \frac{dRT}{P} = \frac{\left(1.33\frac{\text{g}}{\text{L}}\right)\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(150+273)\text{K}}{\left(764\text{ mmHg}\times\frac{1\text{ atm}}{760\text{ mmHg}}\right)} = 45.9\text{ g/mol}$$

Some nitrogen oxides and their molar masses are:

NO	30 g/mol	N_2O	44 g/mol	NO_2	46 g/mol

The nitrogen oxide is most likely $\mathbf{NO_2}$, although N_2O cannot be completely ruled out.

5.135 Assuming a volume of 1.00 L, we can determine the average molar mass of the mixture.

$$n_{\text{mixture}} = \frac{PV}{RT} = \frac{(0.98\text{ atm})(1.00\text{ L})}{\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(25 + 273)\text{K}} = 0.0401\text{ mol}$$

The average molar mass of the mixture is:

$$\overline{\mathcal{M}} = \frac{2.7\text{ g}}{0.0401\text{ mol}} = \mathbf{67\ g/mol}$$

Now, we can calculate the mole fraction of each component of the mixture. Once we determine the mole fractions, we can calculate the partial pressure of each gas.

$$X_{NO_2}\mathcal{M}_{NO_2} + X_{N_2O_4}\mathcal{M}_{N_2O_4} = 67\text{ g/mol}$$

$$X_{NO_2}(46.01\text{ g/mol}) + (1 - X_{NO_2})(92.02\text{ g/mol}) = 67\text{ g/mol}$$

$$46.01X_{NO_2} - 92.02X_{NO_2} + 92.02 = 67$$

$$X_{NO_2} = 0.54 \qquad \text{and} \qquad X_{N_2O_4} = 1 - 0.54 = 0.46$$

Finally, the partial pressures are:

$$P_{NO_2} = X_{NO_2}P_T \qquad P_{N_2O_4} = X_{N_2O_4}P_T$$

$$\boldsymbol{P_{NO_2}} = (0.54)(0.98\text{ atm}) = \mathbf{0.53\ atm} \qquad \boldsymbol{P_{N_2O_4}} = (0.46)(0.98\text{ atm}) = \mathbf{0.45\ atm}$$

5.136 When calculating root-mean-square speed, remember that the molar mass must be in units of kg/mol.

$$\boldsymbol{u_{\text{rms}}} = \sqrt{\frac{3RT}{\mathcal{M}}} = \sqrt{\frac{3(8.314\text{ J/mol}\cdot\text{K})(1.7\times10^{-7}\text{ K})}{85.47\times10^{-3}\text{ kg/mol}}} = \mathbf{7.0\times10^{-3}\ m/s}$$

The mass of one Rb atom in kg is:

$$\frac{85.47\text{ g Rb}}{1\text{ mol Rb}} \times \frac{1\text{ mol Rb}}{6.022\times10^{23}\text{ Rb atoms}} \times \frac{1\text{ kg}}{1000\text{ g}} = 1.419\times10^{-25}\text{ kg/Rb atom}$$

$$\overline{\mathbf{KE}} = \frac{1}{2}m\overline{u^2} = \frac{1}{2}(1.419\times10^{-25}\text{ kg})(7.0\times10^{-3}\text{ m/s})^2 = \mathbf{3.5\times10^{-30}\ J}$$

5.137 First, calculate the moles of hydrogen gas needed to fill a 4.1 L life belt.

$$n_{H_2} = \frac{PV}{RT} = \frac{(0.97\text{ atm})(4.1\text{ L})}{\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(12 + 273)\text{K}} = 0.17\text{ mol }H_2$$

The balanced equation shows a mole ratio between H_2 and LiH of 1:1. Therefore, 0.17 mole of LiH is needed. Converting to mass in grams:

$$\textbf{? g LiH} = 0.17\text{ mol LiH} \times \frac{7.949\text{ g LiH}}{1\text{ mol LiH}} = \textbf{1.4 g LiH}$$

5.138 The molar volume is the volume of 1 mole of gas under the specified conditions.

$$V = \frac{nRT}{P} = \frac{(1\text{ mol})\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(220\text{ K})}{\left(6.0\text{ mmHg} \times \frac{1\text{ atm}}{760\text{ mmHg}}\right)} = \mathbf{2.3 \times 10^3\ L}$$

5.139 $P_{CO_2} = (0.965) \times (9.0 \times 10^6\text{ Pa}) = \mathbf{8.7 \times 10^6\ Pa}$

$P_{N_2} = (0.035) \times (9.0 \times 10^6\text{ Pa}) = \mathbf{3.2 \times 10^5\ Pa}$

$P_{SO_2} = (1.5 \times 10^{-4}) \times (9.0 \times 10^6\text{ Pa}) = \mathbf{1.4 \times 10^3\ Pa}$

5.140 The volume of the bulb can be calculated using the ideal gas equation. Pressure and temperature are given in the problem. Moles of air must be calculated before the volume can be determined.

Mass of air = 91.6843 g – 91.4715 g = 0.2128 g air

Molar mass of air = (0.78 × 28.02 g/mol) + (0.21 × 32.00 g/mol) + (0.01 × 39.95 g/mol) = 29 g/mol

$$\text{moles air} = 0.2128\text{ g air} \times \frac{1\text{ mol air}}{29\text{ g air}} = 7.3 \times 10^{-3}\text{ mol air}$$

Now, we can calculate the volume of the bulb.

$$V_{\textbf{bulb}} = \frac{nRT}{P} = \frac{(7.3 \times 10^{-3}\text{ mol})\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(23 + 273)\text{K}}{\left(744\text{ mmHg} \times \frac{1\text{ atm}}{760\text{ mmHg}}\right)} = 0.18\text{ L} = \mathbf{1.8 \times 10^2\ mL}$$

5.141 **(a)** (i) Since the two He samples are at the same temperature, their rms speeds and the average kinetic energies are the same.

(ii) The He atoms in V_1 (smaller volume) collide with the walls more frequently. Since the average kinetic energies are the same, the force exerted in the collisions is the same in both flasks.

(b) (i) The rms speed is greater at the higher temperature, T_2.

(ii) The He atoms at the higher temperature, T_2, collide with the walls with greater frequency and with greater force.

(c) (i) False. The rms speed is greater for the lighter gas, He.

(ii) True. The gases are at the same temperature.

(iii) True. $u_{\text{rms}} = \sqrt{\dfrac{(3)\left(8.314\dfrac{\text{J}}{\text{mol}\cdot\text{K}}\right)(74 + 273)\text{K}}{4.003 \times 10^{-3}\ \text{kg/mol}}} = \mathbf{1.47 \times 10^3\ m/s}$

5.142 In Problem 5.102, the mass of the Earth's atmosphere was determined to be 5.25×10^{18} kg. Assuming that the molar mass of air is 29.0 g/mol, we can calculate the number of molecules in the atmosphere.

(a) $$(5.25 \times 10^{18}\ \text{kg air}) \times \frac{1000\ \text{g}}{1\ \text{kg}} \times \frac{1\ \text{mol air}}{29.0\ \text{g air}} \times \frac{6.022 \times 10^{23}\ \text{molecules air}}{1\ \text{mol air}} = \mathbf{1.09 \times 10^{44}\ molecules}$$

(b) First, calculate the moles of air exhaled in every breath. (500 mL = 0.500 L)

$$n = \frac{PV}{RT} = \frac{(1\ \text{atm})(0.500\ \text{L})}{\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(37 + 273)\text{K}} = 1.96 \times 10^{-2}\ \text{mol air/breath}$$

Next, convert to molecules of air per breath.

$$1.96 \times 10^{-2}\ \text{mol air/breath} \times \frac{6.022 \times 10^{23}\ \text{molecules air}}{1\ \text{mol air}} = \mathbf{1.18 \times 10^{22}\ molecules/breath}$$

(c) $$\frac{1.18 \times 10^{22}\ \text{molecules}}{1\ \text{breath}} \times \frac{12\ \text{breaths}}{1\ \text{min}} \times \frac{60\ \text{min}}{1\ \text{h}} \times \frac{24\ \text{h}}{1\ \text{day}} \times \frac{365\ \text{days}}{1\ \text{yr}} \times 35\ \text{yr} = \mathbf{2.60 \times 10^{30}\ molecules}$$

(d) Fraction of molecules in the atmosphere exhaled by Mozart is:

$$\frac{2.60 \times 10^{30}\ \text{molecules}}{1.09 \times 10^{44}\ \text{molecules}} = \mathbf{2.39 \times 10^{-14}}$$

Or,

$$\frac{1}{2.39 \times 10^{-14}} = 4.18 \times 10^{13}$$

Thus, about 1 molecule of air in every 4×10^{13} molecules was exhaled by Mozart.

In a single breath containing 1.18×10^{22} molecules, we would breathe in on average:

$$(1.18 \times 10^{22}\ \text{molecules}) \times \frac{1\ \text{Mozart air molecule}}{4 \times 10^{13}\ \text{air molecules}} = \mathbf{3 \times 10^8\ molecules\ that\ Mozart\ exhaled}$$

(e) We made the following assumptions:

1. Complete mixing of air in the atmosphere.
2. That no molecules escaped to the outer atmosphere.
3. That no molecules were used up during metabolism, nitrogen fixation, and so on.

5.143 First, let's calculate the root-mean-square speed of N_2 at 25°C.

$$u_{\text{rms}}(\text{N}_2) = \sqrt{\frac{3RT}{\mathcal{M}}} = \sqrt{\frac{3(8.314\ \text{J/mol}\cdot\text{K})(298\ \text{K})}{0.02802\ \text{kg/mol}}} = 515\ \text{m/s}$$

Now, we can calculate the temperature at which He atoms will have this same root-mean-square speed.

$$u_{\text{rms}}(\text{He}) = \sqrt{\frac{3RT}{\mathcal{M}}}$$

$$515\text{ m/s} = \sqrt{\frac{3(8.314\text{ J/mol}\cdot\text{K})T}{0.004003\text{ kg/mol}}}$$

$$\boldsymbol{T} = \mathbf{42.6\ K}$$

5.144 The ideal gas law can be used to calculate the moles of water vapor per liter.

$$\frac{n}{V} = \frac{P}{RT} = \frac{1.0\text{ atm}}{(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(100+273)\text{K}} = 0.033\,\frac{\text{mol}}{\text{L}}$$

We eventually want to find the distance between molecules. Therefore, let's convert moles to molecules, and convert liters to a volume unit that will allow us to get to distance (m^3).

$$\left(\frac{0.033\text{ mol}}{1\text{ L}}\right)\left(\frac{6.022\times10^{23}\text{ molecules}}{1\text{ mol}}\right)\left(\frac{1000\text{ L}}{1\text{ m}^3}\right) = 2.0\times10^{25}\,\frac{\text{molecules}}{\text{m}^3}$$

This is the number of ideal gas molecules in a cube that is 1 meter on each side. Assuming an equal distribution of molecules along the three mutually perpendicular directions defined by the cube, a linear density in one direction may be found:

$$\left(\frac{2.0\times10^{25}\text{ molecules}}{1\text{ m}^3}\right)^{\frac{1}{3}} = 2.7\times10^{8}\,\frac{\text{molecules}}{\text{m}}$$

This is the number of molecules on a line *one* meter in length. The distance between each molecule is given by:

$$\frac{1\text{ m}}{2.70\times10^{8}} = 3.7\times10^{-9}\text{ m} = \mathbf{3.7\ nm}$$

Assuming a water molecule to be a sphere with a diameter of 0.3 nm, the water molecules are separated by over 12 times their diameter: $\frac{3.7\text{ nm}}{0.3\text{ nm}} \approx 12$ times.

A similar calculation is done for liquid water. Starting with density, we convert to molecules per cubic meter.

$$\frac{0.96\text{ g}}{1\text{ cm}^3}\times\frac{1\text{ mol H}_2\text{O}}{18.02\text{ g H}_2\text{O}}\times\frac{6.022\times10^{23}\text{ molecules}}{1\text{ mol H}_2\text{O}}\times\left(\frac{100\text{ cm}}{1\text{ m}}\right)^3 = 3.2\times10^{28}\,\frac{\text{molecules}}{\text{m}^3}$$

This is the number of liquid water molecules in *one* cubic meter. From this point, the calculation is the same as that for water vapor, and the space between molecules is found using the same assumptions.

$$\left(\frac{3.2\times10^{28}\text{ molecules}}{1\text{ m}^3}\right)^{\frac{1}{3}} = 3.2\times10^{9}\,\frac{\text{molecules}}{\text{m}}$$

$$\frac{1\text{ m}}{3.2\times 10^{9}} = 3.1\times 10^{-10}\text{ m} = \mathbf{0.31\ nm}$$

Assuming a water molecule to be a sphere with a diameter of 0.3 nm, to one significant figure, the water molecules are touching each other in the liquid phase.

5.145 **Radon**, because it is radioactive so that its mass is constantly changing (decreasing). The number of radon atoms is not constant.

5.146 Since the $R = 8.314$ J/mol·K and $1\text{ J} = 1\,\dfrac{\text{kg}\cdot\text{m}^2}{\text{s}^2}$, then the mass substituted into the equation must have units of kg and the height must have units of meters.

29 g/mol = 0.029 kg/mol
5.0 km = 5.0×10^3 m

Substituting the given quantities into the equation, we find the atmospheric pressure at 5.0 km to be:

$$P = P_0 e^{-\frac{g\mathcal{M}h}{RT}}$$

$$P = (1.0\text{ atm})e^{-\left(\frac{(9.8\text{ m/s}^2)(0.029\text{ kg/mol})(5.0\times 10^3\text{ m})}{(8.314\text{ J/mol}\cdot\text{K})(278\text{ K})}\right)}$$

$$\boldsymbol{P} = \mathbf{0.54\ atm}$$

5.147 We need to find the total moles of gas present in the flask after the reaction. Then, we can use the ideal gas equation to calculate the total pressure inside the flask. Since we are given the amounts of both reactants, we need to find which reactant is used up first. This is a limiting reagent problem. Let's calculate the moles of each reactant present.

$$5.72\text{ g C}\times\frac{1\text{ mol C}}{12.01\text{ g C}} = 0.476\text{ mol C}$$

$$68.4\text{ g O}_2\times\frac{1\text{ mol O}_2}{32.00\text{ g O}_2} = 2.14\text{ mol O}_2$$

The mole ratio between C and O_2 in the balanced equation is 1:1. Therefore, C is the limiting reagent. The amount of C remaining after complete reaction is 0 moles. Since the mole ratio between C and O_2 is 1:1, the amount of O_2 that reacts is 0.476 mole. The amount of O_2 remaining after reaction is:

$$\text{moles O}_2\text{ remaining} = \text{moles O}_2\text{ initial} - \text{moles O}_2\text{ reacted} = 2.14\text{ mol} - 0.476\text{ mol} = 1.66\text{ mol O}_2$$

The amount of CO_2 produced in the reaction is :

$$0.476\text{ mol C}\times\frac{1\text{ mol CO}_2}{1\text{ mol C}} = 0.476\text{ mol CO}_2$$

The total moles of gas present after reaction are:

$$\text{total mol of gas} = \text{mol CO}_2 + \text{mol O}_2 = 0.476\text{ mol} + 1.66\text{ mol} = 2.14\text{ mol}$$

Using the ideal gas equation, we can now calculate the total pressure inside the flask.

$$\boldsymbol{P} = \frac{nRT}{V} = \frac{(2.14\text{ mol})(0.0821\text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K})(182 + 273)\text{K}}{8.00\text{ L}} = \mathbf{9.99\ atm}$$

5.148 The reaction between Zn and HCl is: $\text{Zn}(s) + 2\text{HCl}(aq) \rightarrow \text{H}_2(g) + \text{ZnCl}_2(aq)$

From the amount of $\text{H}_2(g)$ produced, we can determine the amount of Zn reacted. Then, using the original mass of the sample, we can calculate the mass % of Zn in the sample.

$$n_{\text{H}_2} = \frac{PV_{\text{H}_2}}{RT}$$

$$n_{\text{H}_2} = \frac{\left(728\text{ mmHg} \times \dfrac{1\text{ atm}}{760\text{ mmHg}}\right)(1.26\text{ L})}{\left(0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(22 + 273)\text{K}} = 0.0498\text{ mol H}_2$$

Since the mole ratio between H_2 and Zn is 1:1, the amount of Zn reacted is also 0.0498 mole. Converting to grams of Zn, we find:

$$0.0498\text{ mol Zn} \times \frac{65.39\text{ g Zn}}{1\text{ mol Zn}} = 3.26\text{ g Zn}$$

The mass percent of Zn in the 6.11 g sample is:

$$\textbf{mass \% Zn} = \frac{\text{mass Zn}}{\text{mass sample}} \times 100\% = \frac{3.26\text{ g}}{6.11\text{ g}} \times 100\% = \mathbf{53.4\%}$$

5.149 When the drum is dented, the volume decreases and therefore we expect the pressure to increase. However, the pressure due to acetone vapor (400 mmHg) will not change as long as the temperature stays at 18°C (vapor pressure is constant at a given temperature). As the pressure increases, more acetone vapor will condense to liquid. Assuming that air does not dissolve in acetone, the pressure inside the drum will increase due an increase in the pressure due to air. Initially, the total pressure is 750 mmHg. The pressure due to air initially is:

$$P_{\text{T}} = P_{\text{air}} + P_{\text{acetone}}$$

$$P_{\text{air}} = P_{\text{T}} - P_{\text{acetone}} = 750\text{ mmHg} - 400\text{ mmHg} = 350\text{ mmHg}$$

The initial volume of vapor in the drum is:

$$V_1 = 25.0\text{ gal} - 15.4\text{ gal} = 9.6\text{ gal}$$

When the drum is dented, the volume the vapor occupies decreases to:

$$V_2 = 20.4\text{ gal} - 15.4\text{ gal} = 5.0\text{ gal}$$

The same number of air molecules now occupies a smaller volume. The pressure increases according to Boyle's Law.

$$P_1V_1 = P_2V_2$$

$$(350\text{ mmHg})(9.6\text{ gal}) = P_2(5.0\text{ gal})$$

$$P_2 = 672\text{ mmHg}$$

This is the pressure due to air. The pressure due to acetone vapor is still 400 mmHg. The total pressure inside the drum after the accident is:

$$P_T = P_{\text{air}} + P_{\text{acetone}}$$

$$\boldsymbol{P_T} = 672\text{ mmHg} + 400\text{ mmHg} = \mathbf{1072\text{ mmHg}}$$

5.150 We start with Graham's Law as this problem relates to effusion of gases. Using Graham's Law, we can calculate the effective molar mass of the mixture of CO and CO_2. Once the effective molar mass of the mixture is known, we can determine the mole fraction of each component. Because $n \propto V$ at constant T and P, the volume fraction = mole fraction.

$$\frac{r_{\text{He}}}{r_{\text{mix}}} = \sqrt{\frac{\mathcal{M}_{\text{mix}}}{\mathcal{M}_{\text{He}}}}$$

$$\mathcal{M}_{\text{mix}} = \left(\frac{r_{\text{He}}}{r_{\text{mix}}}\right)^2 \mathcal{M}_{\text{He}}$$

$$\mathcal{M}_{\text{mix}} = \left(\frac{\dfrac{29.7\text{ mL}}{2.00\text{ min}}}{\dfrac{10.0\text{ mL}}{2.00\text{ min}}}\right)^2 (4.003\text{ g/mol}) = 35.31\text{ g/mol}$$

Now that we know the molar mass of the mixture, we can calculate the mole fraction of each component.

$$X_{\text{CO}} + X_{\text{CO}_2} = 1$$

and

$$X_{\text{CO}_2} = 1 - X_{\text{CO}}$$

The mole fraction of each component multiplied by its molar mass will give the contribution of that component to the effective molar mass.

$$X_{\text{CO}}\mathcal{M}_{\text{CO}} + X_{\text{CO}_2}\mathcal{M}_{\text{CO}_2} = \mathcal{M}_{\text{mix}}$$

$$X_{\text{CO}}\mathcal{M}_{\text{CO}} + (1 - X_{\text{CO}})\mathcal{M}_{\text{CO}_2} = \mathcal{M}_{\text{mix}}$$

$$X_{\text{CO}}(28.01\text{ g/mol}) + (1 - X_{\text{CO}})(44.01\text{ g/mol}) = 35.31\text{ g/mol}$$

$$28.01X_{\text{CO}} + 44.01 - 44.01X_{\text{CO}} = 35.31$$

$$16.00X_{\text{CO}} = 8.70$$

$$X_{\text{CO}} = 0.544$$

At constant P and T, $n \propto V$. Therefore, volume fraction = mole fraction. As a result,

% of CO by volume = **54.4%**

% of CO_2 by volume = 1 – % of CO by volume = **45.6%**

5.151 **(a)** The plots dip due to intermolecular attractions between gas particles. Consider the approach of a particular molecule toward the wall of a container. The intermolecular attractions exerted by its neighbors tend to soften the impact made by this molecule against the wall. The overall effect is a lower gas pressure than we would expect for an ideal gas. Thus, PV/RT decreases. The plots rise because at higher pressures (smaller volumes), the molecules are close together and repulsive forces among them become dominant. Repulsive forces increase the force of impact of the gas molecules with the walls of the container. The overall effect is a greater gas pressure than we would expect for an ideal gas. Hence, $PV/RT > 1$ and the curves rise above the horizontal line.

(b) For 1 mole of an ideal gas, $PV/RT = 1$, no matter what the pressure of the gas. At very low pressures, all gases behave ideally; therefore, PV/RT converges to 1 as the pressure approaches zero. As the pressure of a gas approaches zero, a gas behaves more and more like an ideal gas.

(c) The intercept on the ideal gas line means that $PV/RT = 1$. However, this does ***not*** mean that the gas behaves ideally. It just means that at this particular pressure molecular attraction is equal to molecular repulsion so the net interaction is zero.

5.152 The reactions are:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

For a given volume and temperature, $n \alpha P$. This means that the greater the pressure of reactant, the more moles of reactant, and hence the more product (CO_2) that will be produced. The pressure of CO_2 produced comes from both the combustion of methane and ethane. We set up an equation using the mole ratios from the balanced equation to convert to pressure of CO_2.

$$\left(P_{CH_4} \times \frac{1 \text{ mol } CO_2}{1 \text{ mol } CH_4}\right) + \left(P_{C_2H_6} \times \frac{4 \text{ mol } CO_2}{2 \text{ mol } C_2H_6}\right) = 356 \text{ mmHg } CO_2$$

$$(1) \quad P_{CH_4} + 2P_{C_2H_6} = 356 \text{ mmHg}$$

Also,

$$(2) \quad P_{CH_4} + P_{C_2H_6} = 294 \text{ mmHg}$$

Subtracting equation (2) from equation (1) gives:

$$P_{C_2H_6} = 356 - 294 = 62 \text{ mmHg}$$

$$P_{CH_4} = 294 - 62 = 232 \text{ mmHg}$$

Lastly, because $n \alpha P$, we can solve for the mole fraction of each component using partial pressures.

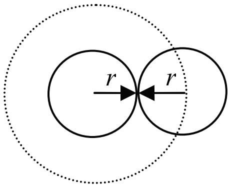
$$\boldsymbol{X_{CH_4}} = \frac{232}{294} = \mathbf{0.789} \qquad \boldsymbol{X_{C_2H_6}} = \frac{62}{294} = \mathbf{0.211}$$

5.153 Consider the motion of an individual molecule. Because of Earth's gravity, the air is more dense near the surface. A "hot" molecule has greater kinetic energy and hence greater speed. It will collide with molecules in all directions. However, because there are more molecules in the downward direction, it experiences more collisions and will soon lose its excess energy. Because the air density decreases with altitude, this molecule will travel farther if it were moving upward. Now, if we start with a large cluster of hot molecules, many of them will rise appreciably before their excess energy dissipates.

5.154 **(a)** We see from the figure that two hard spheres of radius *r* cannot approach each other more closely than 2*r* (measured from the centers). Thus, there is a sphere of radius 2*r* surrounding each hard sphere from which other hard spheres are excluded. The excluded volume/pair of molecules is:

$$V_{\text{excluded/pair}} = \frac{4}{3}\pi(2r)^3 = \frac{32}{3}\pi r^3 = \mathbf{8\left(\frac{4}{3}\pi r^3\right)}$$

This is eight times the volume of an individual molecule.



(b) The result in part (a) is for a pair of molecules, so the excluded volume/molecule is:

$$V_{\text{excluded/molecule}} = \frac{1}{2}\left(\frac{32}{3}\pi r^3\right) = \frac{16}{3}\pi r^3$$

To convert from excluded volume per molecule to excluded volume per mole, we need to multiply by Avogadro's number, N_A.

$$V_{\text{excluded/mole}} = \mathbf{\frac{16}{3}N_A\pi r^3}$$

The sum of the volumes of a mole of molecules (treated as hard spheres of radius *r*) is $\frac{4}{3}N_A\pi r^3$. The excluded volume is **four times** the volume of the spheres themselves.

5.155 Let's calculate the pressure of ammonia using both the ideal gas equation and van der Waal's equation. We can then calculate the percent error in using the ideal gas equation.

Ideal gas: $$P = \frac{nRT}{V} = \frac{(5.00\text{ mol})(0.0821\text{ L}\cdot\text{atm/mol}\cdot\text{K})(300\text{ K})}{1.92\text{ L}} = 64.1\text{ atm}$$

van der Waals: $$P = \frac{nRT}{(V - nb)} - \frac{an^2}{V^2}$$

$$P = \frac{(5.00\text{ mol})(0.0821\text{ L}\cdot\text{atm/mol}\cdot\text{K})(300\text{ K})}{1.92\text{ L} - (5.00\text{ mol})(0.0371\text{ L/mol})} - \frac{(4.17\text{ atm}\cdot\text{L}^2/\text{mol}^2)(5.00\text{ mol})^2}{(1.92\text{ L})^2}$$

$$P = 71.0\text{ atm} - 28.3\text{ atm} = 42.7\text{ atm}$$

The *a* and *b* values used are from Table 5.4 of the text.

The percent error in using the ideal gas equation is:

$$\%\text{ error} = \frac{|P_{\text{vdw}} - P_{\text{ideal}}|}{P_{\text{vdw}}} \times 100\%$$

$$\%\text{ error} = \frac{|42.7\text{ atm} - 64.1\text{ atm}|}{42.7\text{ atm}} \times 100\% = \mathbf{50.1\%}$$

5.156 From the root-mean-square speed, we can calculate the molar mass of the gaseous oxide.

$$u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

$$\mathcal{M} = \frac{3RT}{(u_{\text{rms}})^2} = \frac{3(8.314\ \text{J/mol}\cdot\text{K})(293\ \text{K})}{(493\ \text{m/s})^2} = 0.0301\ \text{kg/mol} = 30.1\ \text{g/mol}$$

The compound must be a monoxide because 2 moles of oxygen atoms would have a mass of 32.00 g. The molar mass of the other element is:

$$30.1\ \text{g/mol} - 16.00\ \text{g/mol} = 14.01\ \text{g/mol}$$

The compound is nitrogen monoxide, **NO**.

5.157 **(a)** The equation to calculate the root-mean-square speed is Equation (5.16) of the text. Let's calculate u_{mp} and u_{rms} at 25°C (298 K). Recall that the molar mass of N_2 must be in units of kg/mol, because the units of R are J/mol·K and 1 J = 1 kg·m^2/s^2.

$$u_{\text{mp}} = \sqrt{\frac{2RT}{\mathcal{M}}} \qquad u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

$$u_{\text{mp}} = \sqrt{\frac{2(8.314\ \text{J/mol}\cdot\text{K})(298\ \text{K})}{0.02802\ \text{kg/mol}}} \qquad u_{\text{rms}} = \sqrt{\frac{3(8.314\ \text{J/mol}\cdot\text{K})(298\ \text{K})}{0.02802\ \text{kg/mol}}}$$

$$\boldsymbol{u_{\text{mp}} = 421\ \text{m/s}} \qquad \boldsymbol{u_{\text{rms}} = 515\ \text{m/s}}$$

The most probable speed (u_{mp}) will always be slower than the root-mean-square speed. We can derive a general relation between the two speeds.

$$\frac{u_{\text{mp}}}{u_{\text{rms}}} = \frac{\sqrt{\frac{2RT}{\mathcal{M}}}}{\sqrt{\frac{3RT}{\mathcal{M}}}} = \sqrt{\frac{\frac{2RT}{\mathcal{M}}}{\frac{3RT}{\mathcal{M}}}}$$

$$\frac{u_{\text{mp}}}{u_{\text{rms}}} = \sqrt{\frac{2}{3}} = 0.816$$

This relation indicates that the most probable speed (u_{mp}) will be 81.6% of the root-mean-square speed (u_{rms}) at a given temperature.

(b) We can derive a relationship between the most probable speeds at T_1 and T_2.

$$\frac{u_{\text{mp}}(1)}{u_{\text{mp}}(2)} = \frac{\sqrt{\frac{2RT_1}{\mathcal{M}}}}{\sqrt{\frac{2RT_2}{\mathcal{M}}}} = \sqrt{\frac{\frac{2RT_1}{\mathcal{M}}}{\frac{2RT_2}{\mathcal{M}}}}$$

$$\frac{u_{\text{mp}}(1)}{u_{\text{mp}}(2)} = \sqrt{\frac{T_1}{T_2}}$$

Looking at the diagram, let's assume that the most probable speed at T_1 = 300 K is 500 m/s, and the most probable speed at T_2 is 1000 m/s. Substitute into the above equation to solve for T_2.

$$\frac{500\text{ m/s}}{1000\text{ m/s}} = \sqrt{\frac{300\text{ K}}{T_2}}$$

$$(0.5)^2 = \frac{300}{T_2}$$

$$\boldsymbol{T_2 = 1200\text{ K}}$$

5.158 Pressure and volume are constant. We start with Equation (5.9) of the text.

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

Because $P_1 = P_2$ and $V_1 = V_2$, this equation reduces to:

$$\frac{1}{n_1T_1} = \frac{1}{n_2T_2}$$

or,

$$n_1T_1 = n_2T_2$$

Because $T_1 = 2T_2$, substituting into the above equation gives:

$$2n_1T_2 = n_2T_2$$

or,

$$2n_1 = n_2$$

This equation indicates that the number of moles of gas after reaction is twice the number of moles of gas before reaction. Only reaction **(b)** fits this description.