CHAPTER 5
GASES

Problem Categories
Biological: 5.38, 5.54, 5.100, 5.132, 5.142.
Descriptive: 5.95, 5.10a, 5.120, 5.125, 5.145.
Environmental: 5.45, 5.53, 5.124, 5.130, 5.133.
Industrial: 5.98, 5.117.

Difficulty Level

5.13 \( \frac{562 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}}{1 \text{ atm}} = 0.739 \text{ atm} \)

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:

\[
\frac{606 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}}{1 \text{ atm}} = 0.797 \text{ atm}
\]

\[
\frac{101.325 \text{ kPa} \times 1 \text{ atm}}{1 \text{ atm}} = 80.8 \text{ kPa}
\]

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 $V \propto \frac{1}{P}$. 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 (2$n$ 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 $V \propto \frac{1}{P}$. 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 $P_1 = 0.970 \text{ atm}$ $P_2 = 0.541 \text{ atm}$ $V_1 = 725 \text{ mL}$ $V_2 =$ ?

$$P_1 V_1 = P_2 V_2$$
$$V_2 = \frac{P_1 V_1}{P_2} = \frac{(0.970 \text{ atm})(725 \text{ mL})}{0.541 \text{ atm}} = 1.30 \times 10^3 \text{ mL}.$$ 

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.

$$\frac{R V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$$

$$P_1 V_1 = P_2 V_2$$

$V_2 = 0.10 V_1$

$$P_2 = \frac{R V_1}{V_2}$$

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

5.21 $P_1 = 1.00 \text{ atm} = 760 \text{ mmHg}$ $P_2 =$ ? $V_1 = 5.80 \text{ L}$ $V_2 = 9.65 \text{ L}$

$$P_1 V_1 = P_2 V_2$$

$$P_2 = \frac{R V_1}{V_2} = \frac{(760 \text{ mmHg})(5.80 \text{ L})}{9.65 \text{ L}} = 457 \text{ mmHg}$$
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.

\[
\frac{R V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}
\]

Because \( n_1 = n_2 \) and \( T_1 = T_2 \),

\[ P_1 V_1 = P_2 V_2 \]

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

<table>
<thead>
<tr>
<th>Initial conditions</th>
<th>Final Conditions</th>
</tr>
</thead>
<tbody>
<tr>
<td>( P_1 = 1.2 \text{ atm} )</td>
<td>( P_2 = 6.6 \text{ atm} )</td>
</tr>
<tr>
<td>( V_1 = 3.8 \text{ L} )</td>
<td>( V_2 = ? )</td>
</tr>
</tbody>
</table>

The final volume is given by:

\[ V_2 = \frac{P_1 V_1}{P_2} \]

\[ V_2 = \frac{(1.2 \text{ atm})(3.8 \text{ L})}{(6.6 \text{ atm})} = 0.69 \text{ L} \]

**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 \( P_1 V_1 = P_2 V_2 \) for this problem. The given information is tabulated below.

<table>
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</tr>
</thead>
<tbody>
<tr>
<td>( P_1 = 1.2 \text{ atm} )</td>
<td>( P_2 = ? )</td>
</tr>
<tr>
<td>( V_1 = 3.8 \text{ L} )</td>
<td>( V_2 = 0.075 \text{ L} )</td>
</tr>
</tbody>
</table>

The final pressure is given by:

\[ P_2 = \frac{P_1 V_1}{V_2} \]

\[ P_2 = \frac{(1.2 \text{ atm})(3.8 \text{ L})}{(0.075 \text{ L})} = 61 \text{ atm} \]

**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 \( T_1 = 25^\circ + 273^\circ = 298 \text{ K} \quad T_2 = 88^\circ + 273^\circ = 361 \text{ K} \)

\( V_1 = 36.4 \text{ L} \quad V_2 = ? \)

\[
\frac{V_1}{T_1} = \frac{V_2}{T_2} \]

\[
V_2 = \frac{V_1 T_2}{T_1} = \frac{36.4 \text{ L}(361 \text{ K})}{298 \text{ K}} = 44.1 \text{ L}
\]

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.

\[
\frac{P V_1}{n_1 T_1} = \frac{P V_2}{n_2 T_2}
\]

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.

<table>
<thead>
<tr>
<th>Initial conditions</th>
<th>Final Conditions</th>
</tr>
</thead>
<tbody>
<tr>
<td>( T_1 = (88 + 273)\text{K} = 361 \text{ K} )</td>
<td>( T_2 = ? )</td>
</tr>
<tr>
<td>( V_1 = 9.6 \text{ L} )</td>
<td>( V_2 = 3.4 \text{ L} )</td>
</tr>
</tbody>
</table>

The final temperature is given by:

\[
T_2 = \frac{T_1 V_2}{V_1} = \frac{(361 \text{ K})(3.4 \text{ L})}{(9.6 \text{ L})} = 1.3 \times 10^2 \text{ K}
\]

5.25 The balanced equation is: \( 4\text{NH}_3(g) + 5\text{O}_2(g) \rightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g) \)

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.
\[ 1\text{Cl}_2(g) + 3\text{F}_2(g) \rightarrow 2\text{Cl}_x\text{F}_3(g) \]

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.

\[ \text{Cl}_2(g) + 3\text{F}_2(g) \rightarrow 2\text{ClF}_3(g) \]

The formula of the product is \( \text{ClF}_3 \).

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

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:

\[ P = \frac{nRT}{V} \]

\[ P = \frac{(6.9 \text{ mof})(0.0821 \left( \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right))(62 + 273)\text{K}}{30.4 \text{ L}} = 6.2 \text{ atm} \]

5.33 We solve the ideal gas equation for \( V \).

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

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.

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<tbody>
<tr>
<td>( T_1 = (25 + 273)\text{K} = 298 \text{ K} )</td>
<td>( T_2 = ? )</td>
</tr>
<tr>
<td>( P_1 = 0.800 \text{ atm} )</td>
<td>( P_2 = 2.00 \text{ atm} )</td>
</tr>
</tbody>
</table>

The final temperature is given by:

\[ T_2 = \frac{T_1P_2}{P_1} \]

\[ T_2 = \frac{(298 \text{ K})(2.00 \text{ atm})}{(0.800 \text{ atm})} = 745 \text{ K} = 472^\circ\text{C} \]
5.35 Initial Conditions
\[ P_1 = 1.2 \text{ atm} \]
\[ V_1 = 2.50 \text{ L} \]
\[ T_1 = (25 + 273) \text{ K} = 298 \text{ K} \]

Final Conditions
\[ P_2 = 3.00 \times 10^{-3} \text{ atm} \]
\[ V_2 = ? \]
\[ T_2 = (-23 + 273) \text{ K} = 250 \text{ K} \]

\[ \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \]

\[ V_2 = \frac{P_1 V_2 T_1}{T_2 P_1} = \frac{(1.2 \text{ atm})(2.50 \text{ L})(250 \text{ K})}{(298 \text{ K})(3.00 \times 10^{-3} \text{ atm})} = 8.4 \times 10^2 \text{ L} \]

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

\[ \frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \]

\[ \frac{P_1}{T_1} = \frac{P_2}{T_2} \]

The given information is tabulated below.

<table>
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<tr>
<td>( T_1 = 273 \text{ K} )</td>
<td>( T_2 = (250 + 273) \text{ K} = 523 \text{ K} )</td>
</tr>
<tr>
<td>( P_1 = 1.0 \text{ atm} )</td>
<td>( P_2 = ? )</td>
</tr>
</tbody>
</table>

The final pressure is given by:

\[ P_2 = \frac{P_1 T_2}{T_1} \]

\[ P_2 = \frac{(1.0 \text{ atm})(523 \text{ K})}{273 \text{ K}} = 1.9 \text{ atm} \]

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 \( \frac{P_2}{P_1} = \frac{1}{3} \).

\[ \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \]

\[ V_2 = V_1 \left( \frac{P_1}{P_2} \right) \left( \frac{T_2}{T_1} \right) = 6.0 \text{ L} \times 3.0 \times 0.50 = 9.0 \text{ L} \]

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

\[ \frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \]

\[ \frac{V_1}{T_1} = \frac{V_2}{T_2} \]
The given information is tabulated below.

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</thead>
<tbody>
<tr>
<td>$T_1 = (20.1 + 273)\ K = 293.1\ K$</td>
<td>$T_2 = (36.5 + 273)\ K = 309.5\ K$</td>
</tr>
<tr>
<td>$V_1 = 0.78\ L$</td>
<td>$V_2 = \ ?$</td>
</tr>
</tbody>
</table>

The final volume is given by:

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

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

5.39 \hspace{1cm} \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}

$$V_1 = \frac{P_2 V_2 T_1}{P_1 T_2} = \frac{(0.60\ atm)(94\ mL)(66 + 273)\ K}{(0.85\ atm)(45 + 273)\ K} = 71\ mL$$

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.

$$?\ mol\ CO_2 = \frac{88.4\ \text{g CO}_2}{44.01\ \text{g CO}_2} \times \frac{1\ mol\ CO_2}{\frac{44.01\ g\ CO_2}{22.41\ L}} = 2.01\ mol\ CO_2$$

We now substitute into the ideal gas equation to calculate volume of CO$_2$.

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

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$_2$.

$$?\ L\ CO_2 = 2.01\ mol\ CO_2 \times \frac{22.41\ L}{1\ mol} = 45.0\ L\ CO_2$$

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 \hspace{1cm} \frac{R V_1}{T_1} = \frac{P_2 V_2}{T_2}

$$V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} = \frac{(772\ mmHg)(6.85\ L)(273\ K)}{(760\ mmHg)(35 + 273)\ K} = 6.17\ L$$
5.42 The molar mass of CO$_2$ = 44.01 g/mol. Since $PV = nRT$, we write:

\[ P = \frac{nRT}{V} \]

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

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

\[ n = \frac{PV}{RT} = \frac{(1.00 \text{ atm})(0.280 \text{ L})}{0.0821 \frac{\text{L atm}}{\text{mol K}}(273 \text{ K})} = 0.0125 \text{ mol} \]

Solving for the molar mass:

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

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.

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

\[ T = 44^\circ + 273^\circ = 317 \text{ K} \]

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

Calculate the molar mass by substituting in the known quantities.

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

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.

\[ n = \frac{PV}{RT} \]
\[ n = \frac{(0.975 \text{ atm})(5.40 \text{ L})}{0.0821 \text{ atm mol}^{-1} \text{ K}^{-1} (317 \text{ K})} = 0.202 \text{ mol} \]

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

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

**5.45** First calculate the moles of ozone (O\(_3\)) using the ideal gas equation.

\[
n = \frac{PV}{RT} = \frac{(1.0 \times 10^{-3} \text{ atm})(1.0 \text{ L})}{0.0821 \text{ L atm}^{-1} \text{ mol}^{-1} \text{ K}^{-1} (250 \text{ K})} = 4.9 \times 10^{-5} \text{ mol O}_3
\]

Use Avogadro’s number as a conversion factor to convert to molecules of O\(_3\).

\[
\text{molecules O}_3 = (4.9 \times 10^{-5} \text{ mol O}_3) \times \frac{6.022 \times 10^{23} \text{ O}_3 \text{ molecules}}{1 \text{ mol O}_3} = 3.0 \times 10^{19} \text{ O}_3 \text{ molecules}
\]

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

\[
\text{Number of particles} = 1.0 \text{ L} \times \frac{1 \text{ mol}}{22.414 \text{ L}} \times \frac{6.022 \times 10^{23} \text{ particles}}{1 \text{ mol}} = 2.7 \times 10^{22} \text{ particles}
\]

\[
\text{Number of N}_2 \text{ molecules} = \left( \frac{78\%}{100\%} \right) (2.7 \times 10^{22} \text{ particles}) = 2.1 \times 10^{22} \text{ N}_2 \text{ molecules}
\]

\[
\text{Number of O}_2 \text{ molecules} = \left( \frac{21\%}{100\%} \right) (2.7 \times 10^{22} \text{ particles}) = 5.7 \times 10^{21} \text{ O}_2 \text{ molecules}
\]

\[
\text{Number of Ar atoms} = \left( \frac{1\%}{100\%} \right) (2.7 \times 10^{22} \text{ particles}) = 3 \times 10^{20} \text{ Ar atoms}
\]

**5.47** The density is given by:

\[
\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{4.65 \text{ g}}{2.10 \text{ L}} = 2.21 \text{ g/L}
\]

Solving for the molar mass:

\[
\text{molar mass} = \frac{dRT}{P} = \frac{(2.21 \text{ g/L}) (0.0821 \text{ L atm}^{-1} \text{ mol}^{-1} \text{ K}^{-1})(27 + 273)\text{ K}}{(1.00 \text{ atm})} = 54.4 \text{ g/mol}
\]

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

\[
d = \frac{P \cdot M}{RT}
\]

\[
M = 1.008 \text{ g/mol} + 79.90 \text{ g/mol} = 80.91 \text{ g/mol}
\]

\[
T = 46^\circ + 273^\circ = 319 \text{ K}
\]

\[
P = 733 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.964 \text{ atm}
\]
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.

\[
V = \frac{nRT}{P} = \frac{(1 \text{ mol})(0.0821 \text{ L·atm/mol·K})(319 \text{ K})}{0.964 \text{ atm}} = 27.2 \text{ L}
\]

Now, we can calculate the density of HBr gas.

\[
\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{80.91 \text{ g}}{27.2 \text{ L}} = 2.97 \text{ g/L}
\]

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\textsubscript{5.40}H\textsubscript{13.4}O\textsubscript{1.35}. Dividing by 1.35 gives the empirical formula, C\textsubscript{4}H\textsubscript{10}O.

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

\[
n = \frac{PV}{RT} = \frac{750 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}}{0.0821 \text{ L·atm/mol·K}(120 + 273)\text{K}} = 0.0306 \text{ mol}
\]

Solving for the molar mass:

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

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, C\textsubscript{4}H\textsubscript{10}O.
METHOD 2:

First calculate the molar mass using the ideal gas equation.

\[
n = \frac{PV}{RT} = \frac{750 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} \times (1.00 \text{ L})}{0.0821 \frac{\text{L} \cdot \text{ atm}}{\text{mol} \cdot \text{K}} \times (120 + 273) \text{K}} = 0.0306 \text{ mol}
\]

Solving for the molar mass:

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

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.

\[
n_C = (0.649) \times (75.2 \text{ g}) \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.06 \text{ mol C}
\]

\[
n_H = (0.135) \times (75.2 \text{ g}) \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 10.07 \text{ mol H}
\]

\[
n_O = (0.216) \times (75.2 \text{ g}) \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.02 \text{ mol O}
\]

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 \( \text{C}_4\text{H}_{10}\text{O} \).

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.

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

\[
T(\text{K}) = 20^\circ + 273^\circ = 293 \text{ K}
\]

\[
\mathcal{M} = \frac{(4.52 \text{ g/L}) \times (0.0821 \frac{\text{L} \cdot \text{ atm}}{\text{mol} \cdot \text{K}} \times (293 \text{ K})}{1.02 \text{ atm}} = 107 \text{ g/mol}
\]

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. \text{ Therefore, the molecular formula is the same as the empirical formula, } \text{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\(_2\).

\[
9.0 \text{ L NO} \times \frac{2 \text{ volumes NO}_2}{2 \text{ volumes NO}} = 9.0 \text{ L NO}_2
\]

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\(_4\) \( \rightarrow \) 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\(_2\).

\[
V = \frac{nRT}{P}
\]

\[
V_{\text{CO}_2} = \frac{(15.0 \text{ mol})(0.0821 \text{ L atm/mol K})(23 + 273)K}{0.985 \text{ atm}} = 3.70 \times 10^2 \text{ L}
\]

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.

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

\[
V = \frac{nRT}{P} = \frac{(79.2 \text{ mol})(0.0821 \text{ L atm/mol K})(303.5K)}{1.12 \text{ atm}} = 1.76 \times 10^3 \text{ L} = 1.76 \times 10^6 \text{ mL SO}_2
\]

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\(_2\).

\[
V = \frac{nRT}{P}
\]
\[ V_{\text{CO}_2} = \frac{(0.0663 \text{ mol})(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(293 \text{ K})}{0.984 \text{ atm}} = 1.62 \text{ L} \]

This is the theoretical yield of CO₂. The actual yield, which is given in the problem, is 1.44 L. We can now calculate the percent yield.

\[
\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100% = \frac{1.44 \text{ L}}{1.62 \text{ L}} \times 100% = 88.9% 
\]

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.

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

Solving for the molar mass:

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

To calculate the empirical formula, first we need to find the mass of F in 0.2631 g of CaF₂.

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

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.

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

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

Thus, we arrive at the formula $\text{P}_{0.003371}\text{F}_{0.006737}$. Dividing by the smallest number of moles (0.003371 mole) gives the empirical formula $\text{PF}_2$.

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 $(\text{PF}_2)_2$ or $\text{P}_2\text{F}_4$. 
**5.56 Strategy:** We can calculate the moles of M reacted, and the moles of H₂ gas produced. By comparing the number of moles of M reacted to the number of moles H₂ produced, we can determine the mole ratio in the balanced equation.

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

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

Solve the ideal gas equation algebraically for \(n_{\text{H}_2}\). Then, calculate the moles of H₂ by substituting the known quantities into the equation.

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

\[
T = 17° + 273° = 290 \text{ K}
\]

\[
\frac{n_{\text{H}_2}}{RT} = \frac{PV_{\text{H}_2}}{RT}
\]

\[
\frac{n_{\text{H}_2}}{RT} = \frac{(0.975 \text{ atm})(0.303 \text{ L})}{0.0821 \frac{\text{L atm}}{\text{mol K}}(290 \text{ K})} = 1.24 \times 10^{-2} \text{ mol H}_2
\]

Compare the number moles of H₂ 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 H₂ to M is 1.5 : 1.

We can now write the balanced equation since we know the mole ratio between H₂ and M. The unbalanced equation is:

\[
\text{M(s) + HCl(aq)} \rightarrow 1.5\text{H}_2(g) + \text{M}_x\text{Cl}_y(aq)
\]

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 MCl₃. The balanced equation is:

\[
\text{M(s) + 3HCl(aq)} \rightarrow 1.5\text{H}_2(g) + \text{MCl}_3(aq)
\]

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 M₂O₃ and M₂(SO₄)₃, respectively.

---

**5.57** The balanced equation for the reaction is: \(\text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s)\)

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

\[
? \text{ mol NH}_3 = 73.0 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} = 4.29 \text{ mol NH}_3
\]
? mol HCl = \( \frac{73.0 \text{ g HCl}}{36.46 \text{ g HCl}} \times 1 \text{ mol HCl} \times 1 \text{ mol HCl} = 2.00 \text{ mol HCl} \)

Since NH\(_3\) and HCl react in a 1:1 mole ratio, HCl is the limiting reagent. The mass of NH\(_4\)Cl formed is:

\[ \text{? g NH}_4\text{Cl} = 2.00 \text{ mol HCl} \times \frac{1 \text{ mol NH}_4\text{Cl}}{1 \text{ mol HCl}} \times \frac{53.49 \text{ g NH}_4\text{Cl}}{1 \text{ mol NH}_4\text{Cl}} = 107 \text{ g NH}_4\text{Cl} \]

The gas remaining is ammonia, NH\(_3\). The number of moles of NH\(_3\) remaining is \((4.29 - 2.00) \text{ mol} = 2.29 \text{ mol NH}_3\). The volume of NH\(_3\) gas is:

\[ V_{\text{NH}_3} = \frac{n_{\text{NH}_3}RT}{P} = \frac{(2.29 \text{ mol}) \left( \frac{0.0821 \text{ L atm}}{\text{mol K}} \right) (14 + 273 \text{ K})}{752 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}} = 54.5 \text{ L NH}_3 \]

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:

\[ \text{CaCO}_3(s) + 2\text{HCl(aq)} \rightarrow \text{CO}_2(g) + \text{CaCl}_2(aq) + \text{H}_2\text{O(l)} \]

The moles of CO\(_2\) produced can be calculated using the ideal gas equation.

\[ n_{\text{CO}_2} = \frac{PV_{\text{CO}_2}}{RT} \]

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

The balanced equation shows a 1:1 mole ratio between CO\(_2\) and CaCO\(_3\). Therefore, \(2.84 \times 10^{-2}\) mole of CaCO\(_3\) must have reacted.

\[ \text{? g CaCO}_3 \text{ reacted} = (2.84 \times 10^{-2} \text{ mol CaCO}_3) \times \frac{100.1 \text{ g CaCO}_3}{1 \text{ mol CaCO}_3} = 2.84 \text{ g CaCO}_3 \]

The percent by mass of the CaCO\(_3\) sample is:

\[ \% \text{ CaCO}_3 = \frac{2.84 \text{ g}}{3.00 \text{ g}} \times 100\% = 94.7\% \]

Assumption: The impurity (or impurities) must not react with HCl to produce CO\(_2\) gas.

5.59 The balanced equation is: \( \text{H}_2(g) + \text{Cl}_2(g) \rightarrow 2\text{HCl(g)} \)

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:

\[
? \text{ g } \text{HCl} = 0.25 \text{ mol H}_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol H}_2} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} = 18 \text{ g HCl}
\]

5.60 The balanced equation is:

\[
\text{C}_2\text{H}_5\text{OH}(l) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(l)
\]

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:

\[
V_{O_2} = \frac{n_{O_2}RT}{P} = \left(14.8 \text{ mol O}_2\right)\left(\frac{0.0821 \text{ L} \cdot \text{ atm}}{\text{mol} \cdot \text{K}}\right)\left(35 + 273 \text{ K}\right) = 3.60 \times 10^2 \text{ L O}_2
\]

Since air is 21.0 percent O\(_2\) by volume, we can write:

\[
V_{\text{air}} = V_{O_2} \left(\frac{100\% \text{ air}}{100\% \text{ O}_2}\right) = \left(3.60 \times 10^2 \text{ L O}_2\right)\left(\frac{100\% \text{ air}}{21\% \text{ O}_2}\right) = 1.71 \times 10^3 \text{ L air}
\]

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:

\[
n = n_{\text{CH}_4} + n_{\text{C}_2\text{H}_6} + n_{\text{C}_3\text{H}_8} = 0.31 \text{ mol} + 0.25 \text{ mol} + 0.29 \text{ mol} = 0.85 \text{ mol}
\]

\[
X_{\text{CH}_4} = \frac{0.31 \text{ mol}}{0.85 \text{ mol}} = 0.36 \quad X_{\text{C}_2\text{H}_6} = \frac{0.25 \text{ mol}}{0.85 \text{ mol}} = 0.29 \quad X_{\text{C}_3\text{H}_8} = \frac{0.29 \text{ mol}}{0.85 \text{ mol}} = 0.34
\]

The partial pressures are:

\[
P_{\text{CH}_4} = X_{\text{CH}_4} \times P_{\text{total}} = 0.36 \times 1.50 \text{ atm} = 0.54 \text{ atm}
\]

\[
P_{\text{C}_2\text{H}_6} = X_{\text{C}_2\text{H}_6} \times P_{\text{total}} = 0.29 \times 1.50 \text{ atm} = 0.44 \text{ atm}
\]

\[
P_{\text{C}_3\text{H}_8} = X_{\text{C}_3\text{H}_8} \times P_{\text{total}} = 0.34 \times 1.50 \text{ atm} = 0.51 \text{ atm}
\]

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

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

(b) We know:

<table>
<thead>
<tr>
<th>Initial conditions</th>
<th>Final Conditions</th>
</tr>
</thead>
<tbody>
<tr>
<td>(P_1 = (0.15 + 0.42)\text{ atm} = 0.57 \text{ atm})</td>
<td>(P_2 = 1.0 \text{ atm})</td>
</tr>
<tr>
<td>(T_1 = (15 + 273)\text{ K} = 288 \text{ K})</td>
<td>(T_2 = 273 \text{ K})</td>
</tr>
<tr>
<td>(V_1 = 2.5 \text{ L})</td>
<td>(V_2 = ?)</td>
</tr>
</tbody>
</table>
Because \( n_1 = n_2 \), we can write:

\[
V_2 = \frac{PV_1 T_2}{P_1 T_1}
\]

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

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

\[
X_{\text{N}_2} = 0.7808 \quad X_{\text{O}_2} = 0.2094 \quad X_{\text{Ar}} = 0.0093 \quad X_{\text{CO}_2} = 0.0005
\]

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

\[
P_{\text{N}_2} = 0.781 \text{ atm}, \quad P_{\text{O}_2} = 0.209 \text{ atm}, \quad P_{\text{Ar}} = 9.3 \times 10^{-3} \text{ atm}, \quad P_{\text{CO}_2} = 5 \times 10^{-4} \text{ atm}
\]

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

\[
c_{\text{N}_2} = \frac{0.781 \text{ atm}}{0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}(273 \text{ K})} = 3.48 \times 10^{-2} \text{ M}
\]

Similarly, \( c_{\text{O}_2} = 9.32 \times 10^{-3} \text{ M}, \quad c_{\text{Ar}} = 4.1 \times 10^{-4} \text{ M}, \quad c_{\text{CO}_2} = 2 \times 10^{-5} \text{ M} \)

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

In this case,

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

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

\[
P_{\text{Ne}} = 745 \text{ mm Hg} - 368 \text{ mmHg} - 28.3 \text{ mmHg} = 349 \text{ mmHg}
\]

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})}{(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(25 + 273)\text{K}} = 0.0098 \text{ mol H}_2
The balanced equation shows a 2:1 mole ratio between Na and H₂. The mass of Na consumed in the reaction is:

\[ ? \text{ g Na} = 0.0098 \text{ mol H}_2 \times \frac{2 \text{ mol Na}}{1 \text{ mol H}_2} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 0.45 \text{ g Na} \]

5.68 **Strategy:** To solve for moles of H₂ generated, we must first calculate the partial pressure of H₂ in the mixture. What gas law do we need? How do we convert from moles of H₂ 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,

\[ P_{\text{Total}} = P_{\text{H}_2} + P_{\text{H}_2O} \]
\[ P_{\text{H}_2} = P_{\text{Total}} - P_{\text{H}_2O} \]
\[ P_{\text{H}_2} = 0.980 \text{ atm} - (23.8 \text{ mmHg}) \left( \frac{1 \text{ atm}}{760 \text{ mmHg}} \right) = 0.949 \text{ atm} \]

Now that we know the pressure of H₂ gas, we can calculate the moles of H₂. Then, using the mole ratio from the balanced equation, we can calculate moles of Zn.

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

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.

\[ ? \text{ g Zn} = 0.303 \text{ mol H}_2 \times \frac{1 \text{ mol Zn}}{1 \text{ mol H}_2} \times \frac{65.39 \text{ g Zn}}{1 \text{ mol Zn}} = 19.8 \text{ g Zn} \]

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 \( P_1 = X_i P_T \). The mole fraction of oxygen is:

\[ X_{\text{O}_2} = \frac{P_{\text{O}_2}}{P_{\text{total}}} = \frac{0.20 \text{ atm}}{4.2 \text{ atm}} = 0.048 \]

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

5.70 \( P_1 = X_i P_T \)

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.

\[ 2\text{NH}_3(g) \longrightarrow \text{N}_2(g) + 3\text{H}_2(g) \]

The mole fractions of H₂ and N₂ are:

\[ X_{\text{H}_2} = \frac{3 \text{ mol}}{3 \text{ mol} + 1 \text{ mol}} = 0.750 \]
The partial pressures of \( \text{H}_2 \) and \( \text{N}_2 \) are:

\[
P_{\text{H}_2} = X_{\text{H}_2} P_T = (0.750)(866 \text{ mmHg}) = 650 \text{ mmHg}
\]

\[
P_{\text{N}_2} = X_{\text{N}_2} P_T = (0.250)(866 \text{ mmHg}) = 217 \text{ mmHg}
\]

5.71 (a) The mole fraction of A is:

\[
X_A = \frac{\text{mol A}}{\text{total mol}}
\]

Because number of particles is directly proportional to number of moles, we can plug-in the number of particles of A and the total number of particles into the above equation to solve for the mole fraction of A.

(i) \( X_A = \frac{4}{9} = 0.44 \)

(ii) \( X_A = \frac{5}{12} = 0.42 \)

(iii) \( X_A = \frac{6}{15} = 0.40 \)

The container with the smallest mole fraction of A is container (iii). Note that even though container (iii) contains the most A particles, the mole fraction of A is the smallest because there is greater total number of particles.

(b) The partial pressure of a gas is proportional to its mole fraction and the total pressure.

\[
P_i = X_i P_T
\]

Therefore, we can calculate the partial pressure of B if we know the mole fraction of B and the total pressure in each container. Because the containers all have the same volume and the same temperature, the pressure in each container will be directly proportional to the moles of gas (or particles of gas). We cannot calculate the pressure directly because the volume and temperature are not known, but the pressures will be proportional to the number of particles in each container (9, 12, and 15, respectively). Let’s assume that the pressures are in units of atmospheres: (i) 9 atm, (ii) 12 atm, (iii) 15 atm.

The mole fraction of B is:

\[
X_B = \frac{\text{mol B}}{\text{total mol}}
\]

Because number of particles is directly proportional to number of moles, we can plug-in the number of particles of B and the total number of particles into the above equation to solve for the mole fraction of B.

(i) \( X_B = \frac{2}{9} = 0.22 \)

(ii) \( X_B = \frac{3}{12} = 0.25 \)

(iii) \( X_B = \frac{4}{15} = 0.27 \)
Now, the partial pressure of B in each container can be calculated \( P_i = X_i P_T \).

(i) \( P_B = (0.22)(9 \text{ atm}) = 2.0 \text{ atm} \)
(ii) \( P_B = (0.25)(12 \text{ atm}) = 3.0 \text{ atm} \)
(iii) \( P_B = (0.27)(15 \text{ atm}) = 4.1 \text{ atm} \)

The container with the highest partial pressure of B is container (iii). Note that once it was known that container (iii) had the greatest mole fraction of B of the three containers and the highest total pressure, we could have reached the same conclusion that container (iii) has the highest partial pressure of B without calculating the partial pressure of B.

\[5.72\] (a) There are 9 particles total in the left container, and 20 particles total in the right container. Because the volume of the right container is double the left, twice as many particles would be needed in the right container to have the same pressure as the left container at the same temperature. We have more than twice the number of particles in the right container compared to the left; hence, the **right container** will have the greater total pressure.

(b) The partial pressure of a gas is proportional to its mole fraction and the total pressure.

\[ P_i = X_i P_T \]

The mole fraction of He is:

\[ X_{\text{He}} = \frac{\text{mol He}}{\text{total mol}} \]

Because number of particles is directly proportional to number of moles, we can plug-in the number of particles of He and the total number of particles into the above equation to solve for the mole fraction of He.

(left) \( X_{\text{He}} = \frac{4}{9} = 0.44 \)

(right) \( X_{\text{He}} = \frac{9}{20} = 0.45 \)

Because the **left container** has the smaller mole fraction of He and the smaller total pressure, then it will have the lower partial pressure of helium \( (P_{\text{He}} = X_{\text{He}} P_T) \).

5.77 \( u_{\text{rms}} = \sqrt{\frac{3RT}{HM}} \)

\[ \text{O}_2: \quad u_{\text{rms}} = \sqrt{\frac{3(8.314 \text{ J/K\cdotmol})(65 + 273)\text{K}}{32.00 \times 10^{-3} \text{ kg/mol}}} = 513 \text{ m/s} \]

\[ \text{UF}_6: \quad u_{\text{rms}} = \sqrt{\frac{3(8.314 \text{ J/K\cdotmol})(65 + 273)\text{K}}{352.00 \times 10^{-3} \text{ kg/mol}}} = 155 \text{ m/s} \]

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 \( M \) so the \( u_{\text{rms}} \) will be expressed in units of m/s.

**Solution:** To calculate \( u_{\text{rms}} \), the units of \( R \) should be 8.314 J/mol\cdotK, and because 1 J = 1 kg\( \cdot \)m\(^2\)/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.

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

\[
u_{\text{rms}}(N_2) = \sqrt{\frac{(3)(8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}})\left(-23 + 273\right)\text{K}}{0.02802 \frac{\text{kg}}{\text{mol}}}} = 472 \text{ m/s}
\]

Similarly,

\[
u_{\text{rms}}(O_2) = 441 \text{ m/s} \quad \nu_{\text{rms}}(O_3) = 360 \text{ m/s}
\]

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

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

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 ($u_{\text{rms}}$) of a gas can be calculated as follows:

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

The rate of diffusion ($\eta$) 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:

\[
\eta_1 = \sqrt{\frac{3RT}{\mathcal{M}_1}} \quad \eta_2 = \sqrt{\frac{3RT}{\mathcal{M}_2}}
\]
Dividing $r_1$ by $r_2$ gives:

$$\frac{r_1}{r_2} = \sqrt{\frac{3RT}{M_1}} \sqrt{\frac{3RT}{M_2}}$$

Canceling $3RT$ from the equation gives:

$$\frac{r_1}{r_2} = \sqrt{\frac{1}{M_1}} \sqrt{\frac{M_2}{M_1}}$$

Does the derived equation make sense? Assume that gas1 is a lighter gas (has a smaller molar mass) than gas2. Dividing a larger molar mass ($M_2$) by a smaller molar mass ($M_1$) 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:

$$s = \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

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:

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

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

$$\frac{r_1}{r_2} = \frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}}$$

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

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

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

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

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)$_x$.

\[
\frac{n_1}{n_2} = \sqrt{\frac{M_2}{M_1}}
\]

\[
\frac{3.3}{1.0} = \sqrt{\frac{M_{\text{Ni(CO)}_x}}{16.04 \text{ g/mol}}}
\]

\[
10.89 = \frac{M_{\text{Ni(CO)}_x}}{16.04 \text{ g/mol}}
\]

\[
M_{\text{Ni(CO)}_x} = 174.7 \text{ g/mol}
\]

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.

\[
\frac{116.0 \text{ g}}{28.01 \text{ g}} = 4.141 \approx 4
\]

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.

van der waals’ equation:

We find the pressure by first solving algebraically for $P$.

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

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

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

ideal gas equation:

\[
P = \frac{nRT}{V} = \frac{(2.50 \text{ mol})(0.0821 \text{ atm} \cdot \text{L/mol} \cdot \text{K})(450 \text{ K})}{(5.00 \text{ L})} = 18.5 \text{ atm}
\]

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.

\[
T(K) = 27°C + 273° = 300 K
\]

\[
P = \frac{nRT}{V} = \frac{(10.0 \text{ mol})(0.0821 \frac{L\cdot atm}{mol\cdot K})(300 \text{ K})}{1.50 L} = 164 \text{ atm}
\]

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

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

Calculating the volume of products:

\[
V_{\text{product}} = \frac{n_{\text{product}}RT}{P} = \frac{(8.3 \text{ mol})(0.0821 \frac{L\cdot atm}{mol\cdot K})(298 K)}{(1.2 \text{ atm})} = 1.7 \times 10^2 \text{ L}
\]

The relationship between partial pressure and \(P_{\text{total}}\) is:

\[
P_i = X_iP_T
\]

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 \(X_{\text{O}_2} = 0.034\).
\[ P_{\text{CO}_2} = X_{\text{CO}_2} P_1 \]

\[ P_{\text{CO}_2} = (0.41)(1.2 \text{ atm}) = 0.49 \text{ atm} \]

Similarly, \( P_{\text{H}_2\text{O}} = 0.41 \text{ atm}, P_{\text{N}_2} = 0.25 \text{ atm}, \) and \( P_{\text{O}_2} = 0.041 \text{ atm}. \)

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.

\[
n = \frac{PV}{RT} = \frac{(0.74 \text{ atm})(97.2 \text{ mL} \times \frac{0.001 \text{ L}}{1 \text{ mL}})}{0.0821 \text{ L atm/mol K}(200 + 273) \text{ K}} = 1.85 \times 10^{-3} \text{ mol}
\]

Molar mass = \( \frac{0.145 \text{ g}}{1.85 \times 10^{-3} \text{ mol}} = 78.4 \text{ g/mol} \)

The empirical mass of CH = 13.02 g/mol

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

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

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

volume \( \text{N}_2 \rightarrow \) moles \( \text{N}_2 \rightarrow \) moles \( \text{NH}_4\text{NO}_2 \rightarrow \) grams \( \text{NH}_4\text{NO}_2 \)

First, calculate the moles of \( \text{N}_2 \) using the ideal gas equation.

\[
T(\text{K}) = 22^\circ + 273^\circ = 295 \text{ K}
\]

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

\[
n_{\text{N}_2} = \frac{P_{\text{N}_2} V}{RT} = \frac{(1.20 \text{ atm})(0.0862 \text{ L})}{0.0821 \text{ L atm/mol K}(295 \text{ K})} = 4.27 \times 10^{-3} \text{ mol}
\]

Next, calculate the mass of \( \text{NH}_4\text{NO}_2 \) needed to produce \( 4.27 \times 10^{-3} \) mole of \( \text{N}_2 \).

\[
\text{g} \ \text{NH}_4\text{NO}_2 = (4.27 \times 10^{-3} \text{ mol} \text{N}_2) \times \frac{1 \text{ mol} \text{NH}_4\text{NO}_2}{1 \text{ mol} \text{N}_2} \times \frac{64.05 \text{ g} \ \text{NH}_4\text{NO}_2}{1 \text{ mol} \text{NH}_4\text{NO}_2} = 0.273 \text{ g}
\]

5.96 The reaction is: \( \text{HCO}_3^-(aq) + \text{H}^+(aq) \rightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g) \)

The mass of \( \text{HCO}_3^- \) reacted is:

\[
3.29 \text{ g tablet} \times \frac{32.5\% \ \text{HCO}_3^-}{100\% \ \text{tablet}} = 1.07 \text{ g HCO}_3^-
\]
\[
\text{mol CO}_2 \text{ produced} = 1.07 \text{ g HCO}_3^- \times \frac{1 \text{ mol HCO}_3^-}{61.02 \text{ g HCO}_3^-} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol HCO}_3^-} = 0.0175 \text{ mol CO}_2
\]

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

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:

\[
P = \frac{nRT}{V} = \frac{(1.47 \text{ mol})(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(43 + 273)\text{K}}{4.00 \text{ L}} = 9.53 \text{ atm}
\]

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

\[
\text{Ni(CO)}_4(g) \longrightarrow \text{Ni(s)} + 4\text{CO(g)}
\]

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.

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

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

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

The most common gases present in exhaled air are: CO\(_2\), O\(_2\), N\(_2\), and H\(_2\)O.

5.101 (a) Write a balanced chemical equation.

\[
2\text{NaHCO}_3(s) \longrightarrow \text{Na}_2\text{CO}_3(s) + \text{CO}_2(g) + \text{H}_2\text{O}(g)
\]

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.

\[ T(K) = 180^\circ + 273^\circ = 453 \text{ K} \]

\[ V_{CO_2} = \frac{n_{CO_2}RT}{P} \]

\[ V_{CO_2} = \frac{(0.030 \text{ mol})(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(453 \text{ K})}{(1.3 \text{ atm})} = 0.86 \text{ L} \]

(b) The balanced chemical equation for the decomposition of NH$_4$HCO$_3$ is

\[ \text{NH}_4\text{HCO}_3(s) \rightarrow \text{NH}_3(g) + \text{CO}_2(g) + \text{H}_2\text{O}(g) \]

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$^2$) \times (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:

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

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

5.103 First, calculate the moles of H$_2$ formed.

\[ ? \text{ mol } H_2 = 3.12 \text{ g} / \text{Al} \times \frac{1 \text{ mol AI}}{26.98 \text{ g}} \times \frac{3 \text{ mol } H_2}{2 \text{ mol AI}} = 0.173 \text{ mol} \]

Next, calculate the volume of H$_2$ produced using the ideal gas equation.

\[ V_{H_2} = \frac{n_{H_2}RT}{P} = \frac{(0.173 \text{ mol})(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(296 \text{ K})}{(1.00 \text{ atm})} = 4.20 \text{ L} \]

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:

\[ \text{NaOH}(aq) + \text{HCl}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaCl}(aq) \]

The number of moles of HCl gas is found from the ideal gas equation. \( V = 0.189 \text{ L}, T = (25 + 273)\text{K} = 298 \text{ K}, \) and \( P = 108 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.142 \text{ atm} \).
\[ n_{\text{HCl}} = \frac{PV_{\text{HCl}}}{RT} = \frac{(0.142 \text{ atm})(0.189 \text{ L})}{(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298 \text{ K})} = 1.10 \times 10^{-3} \text{ mol HCl} \]

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:

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

5.105 (a) \( \text{C}_3\text{H}_8(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g) \)

(b) From the balanced equation, we see that there is a 1:3 mole ratio between \( \text{C}_3\text{H}_8 \) and \( \text{CO}_2 \).

\[ ? \text{ L CO}_2 = 7.45 \text{ g C}_3\text{H}_8 \times \frac{1 \text{ mol C}_3\text{H}_8}{44.09 \text{ g C}_3\text{H}_8} \times \frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} \times \frac{22.414 \text{ L CO}_2}{1 \text{ mol CO}_2} = 11.4 \text{ L CO}_2 \]

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).

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

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

The total pressure is:

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

\[ P_1 = X_1P_T \]. The partial pressures of He and Ne are:

\[ P_{\text{He}} = \frac{0.032 \text{ mol}}{(0.032 + 0.40)\text{mol}} \times 2.2 \text{ atm} = 0.16 \text{ atm} \]

\[ P_{\text{Ne}} = \frac{0.40 \text{ mol}}{(0.032 + 0.40)\text{mol}} \times 2.2 \text{ atm} = 2.0 \text{ atm} \]

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

\[ n_{\text{NO}} = \frac{PV}{RT} = \frac{(0.500 \text{ atm})(4.00 \text{ L})}{(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298 \text{ K})} = 0.0817 \text{ mol NO} \]

\[ n_{\text{O}_2} = \frac{PV}{RT} = \frac{(1.00 \text{ atm})(2.00 \text{ L})}{(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298 \text{ K})} = 0.0817 \text{ mol O}_2 \]
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.

\[
\text{mol NO} = 0 \text{ mol (All NO is consumed during reaction)}
\]

\[
\text{mol NO}_2 = 0.0817 \text{ mol NO} \times \frac{2 \text{ mol NO}_2}{2 \text{ mol NO}} = 0.0817 \text{ mol NO}_2
\]

\[
\text{mol O}_2 \text{ consumed} = 0.0817 \text{ mol NO} \times \frac{1 \text{ mol O}_2}{2 \text{ mol NO}} = 0.0409 \text{ mol O}_2 \text{ consumed}
\]

\[
\text{mol O}_2 \text{ remaining} = 0.0817 \text{ mol O}_2 \text{ initial} - 0.0409 \text{ mol O}_2 \text{ consumed} = 0.0408 \text{ mol O}_2
\]

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

\[
T(K) = 25^\circ + 273^\circ = 298 \text{ K}
\]

\[
P_{O_2} = \frac{n_{O_2}RT}{V} = \frac{(0.0408 \text{ mol})(0.0821 \text{ L atm/mol K})(298 \text{ K})}{(6.00 \text{ L})} = 0.166 \text{ atm}
\]

\[
P_{NO_2} = \frac{n_{NO_2}RT}{V} = \frac{(0.0817 \text{ mol})(0.0821 \text{ L atm/mol K})(298 \text{ K})}{(6.00 \text{ L})} = 0.333 \text{ atm}
\]

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 (\(P_{\text{Total}}\)) 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.

\[
\text{CO}_2(g) + 2\text{NaOH(aq)} \rightarrow \text{Na}_2\text{CO}_3(aq) + \text{H}_2\text{O(l)}
\]

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.

\[
P_{\text{CO}_2} = P_{\text{Total}} - P_{\text{H}_2}
\]

(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.

\[
\frac{n_{\text{H}_2\text{O}}}{RT} = \frac{(24.8 \text{ atm})(2.00 \text{ L})}{0.0821 \text{L atm mol}^{-1} \text{K}^{-1}((120 + 273) \text{K})} = 1.537 \text{ mol H}_2\text{O}
\]

Next, we can determine the mass of H\text{O} in the 54.2 g sample. Subtracting the mass of H\text{O} from 54.2 g will give the mass of MgSO\text{4} in the sample.

\[
1.537 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 27.7 \text{ g H}_2\text{O}
\]

Mass MgSO\text{4} = 54.2 g sample − 27.7 g H\text{O} = 26.5 g MgSO\text{4}

Finally, we can calculate the moles of MgSO\text{4} in the sample. Comparing moles of MgSO\text{4} to moles of H\text{O} will allow us to determine the correct mole ratio in the formula.

\[
\frac{26.5 \text{ g MgSO}_4 \times 1 \text{ mol MgSO}_4}{120.4 \text{ g MgSO}_4} = 0.220 \text{ mol MgSO}_4
\]

\[
\frac{\text{mol H}_2\text{O}}{\text{mol MgSO}_4} = \frac{1.54 \text{ mol}}{0.220 \text{ mol}} = 7.00
\]

Therefore, the mole ratio between H\text{O} and MgSO\text{4} in the compound is 7 : 1. Thus, the value of \(x = 7\), and the formula is MgSO\text{4}\cdot7\text{H}_2\text{O}.

5.111 The reactions are:

\[
\text{Na}_2\text{CO}_3(s) + 2\text{HCl}(aq) \rightarrow 2\text{NaCl}(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g)
\]

\[
\text{MgCO}_3(s) + 2\text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g)
\]

First, let’s calculate the moles of CO\text{2} produced using the ideal gas equation. We carry an extra significant figure throughout this calculation to limit rounding errors.

\[
\frac{n_{\text{CO}_2}}{RT} = \frac{(1.24 \text{ atm})(1.67 \text{ L})}{0.0821 \text{L atm mol}^{-1} \text{K}^{-1}((26 + 273) \text{K})} = 0.08436 \text{ mol CO}_2
\]

Since there is a 1:1 mole ratio between CO\text{2} and both reactants (Na\text{2}CO\text{3} and MgCO\text{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\text{2}CO\text{3} in the mixture, then \((7.63 - x)\) is the mass of MgCO\text{3} in the mixture.

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

\[
0.009434x - 0.01186x + 0.09049 = 0.08436
\]

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

The percent composition by mass of Na\text{2}CO\text{3} in the mixture is:

\[
\text{mass } \% \text{ Na}_2\text{CO}_3 = \frac{\text{mass Na}_2\text{CO}_3}{\text{mass of mixture}} \times 100\% = \frac{2.527 \text{ g}}{7.63 \text{ g}} \times 100\% = 33.1\% \text{ Na}_2\text{CO}_3
\]
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:

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

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

(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 \(4.58 \times 10^{-4} \text{ s}\).

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

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

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 \(\nu_{\text{rms}}\) 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.

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

? g H\textsubscript{2}O vaporized = 0.0222 mol H\textsubscript{2}O \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 0.400 \text{ g H}_2\text{O}

5.114 The moles of O\textsubscript{2} can be calculated from the ideal gas equation. The mass of O\textsubscript{2} can then be calculated using the molar mass as a conversion factor.

\[
n_{\text{O}_2} = \frac{PV}{RT} = \frac{(132 \text{ atm})(120 \text{ L})}{0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1}(22 + 273)\text{K}} = 654 \text{ mol O}_2
\]

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

The volume of O\textsubscript{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\textsubscript{2} = 654 moles.

\[
V_{\text{O}_2} = \frac{n_{\text{O}_2}RT}{P} = \frac{(654 \text{ mol})\left(0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1}\right)(22 + 273)\text{K}}{1.00 \text{ atm}} = 1.58 \times 10^4 \text{ L O}_2
\]
The air inside the egg expands with increasing temperature. The increased pressure can cause the egg to crack.

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

The balanced equation is: \( \text{CO}_2(g) + 2\text{NH}_3(g) \rightarrow (\text{NH}_2)\text{CO(s)} + \text{H}_2\text{O(g)} \)

First, we can calculate the moles of \( \text{NH}_3 \) needed to produce 1.0 ton of urea. Then, we can use the ideal gas equation to calculate the volume of \( \text{NH}_3 \).

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

\[
V_{\text{NH}_3} = \frac{nRT}{P} = \frac{(3.0 \times 10^4 \text{ mol})(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(200 + 273)\text{K}}{150 \text{ atm}} = 7.8 \times 10^3 \text{ L } \text{NH}_3
\]

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.

(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:

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

\[
P_1V_1 = P_2V_2 \quad \text{or} \quad \frac{V_1}{V_2} = \frac{P_2}{P_1}
\]

\[
\frac{V_1}{V_2} = \frac{2.1 \text{ atm}}{1 \text{ atm}} = 2.1
\]

The diver’s lungs would increase in volume \textbf{2.1 times} by the time he reaches the surface.

(b) \( P_{\text{O}_2} = X_{\text{O}_2}P_T \)

\[
P_{\text{O}_2} = \frac{n_{\text{O}_2}}{n_{\text{O}_2} + n_{\text{N}_2}}P_T
\]

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

\[
P_{\text{O}_2} = \frac{V_{\text{O}_2}}{V_{\text{O}_2} + V_{\text{N}_2}}P_T
\]

We know the partial pressure of \( \text{O}_2 \) in air, and we know the total pressure exerted on the diver. Plugging these values into the above equation gives:

\[
0.20 \text{ atm} = \frac{V_{\text{O}_2}}{V_{\text{O}_2} + V_{\text{N}_2}}(4.00 \text{ atm})
\]
\[
\frac{V_{O_2}}{V_{O_2} + V_{N_2}} = \frac{0.20 \text{ atm}}{4.00 \text{ atm}} = 0.050
\]

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

5.120 (a) \[ \text{NH}_4\text{NO}_3(s) \longrightarrow \text{N}_2\text{O}(g) + 2\text{H}_2\text{O}(l) \]

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

5.121 Since Ne and NH\(_3\) are at the same temperature, they have the same average kinetic energy.

\[ KE = \frac{1}{2} mu^2 \]

or

\[ \frac{1}{u^2} = \frac{2KE}{m} \]

Recall that mass must have units of kg because kinetic energy has units of Joules. \[ 1 \text{ J} = 1 \text{ kg} \cdot \text{m}^2 \cdot \text{s}^{-2} \]

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

\[ \frac{20.18 \text{ g Ne}}{1 \text{ mol Ne}} \times \frac{1 \text{ mol Ne}}{6.022 \times 10^{23} \text{ Ne atoms}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 3.351 \times 10^{-26} \text{ kg/Ne atom} \]

Solving for \( \frac{1}{u^2} \):

\[ \frac{1}{u^2} = \frac{2KE}{m} = \frac{2(7.1 \times 10^{-21} \text{ J/atom})}{3.351 \times 10^{-26} \text{ kg/atom}} = 4.2 \times 10^5 \text{ m}^2 \cdot \text{s}^{-2} \]

5.122 The value of \( a \) indicates how strongly molecules of a given type of gas attract one another. \( \text{C}_6\text{H}_6 \) 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.

\[ n = \frac{PV}{RT} = \frac{(1 \text{ atm})(5.58 \times 10^{-3} \text{ L})}{(0.0821 \text{ L atm mol}^{-1} \text{ \cdot K})(45 + 273)\text{K}} = 2.14 \times 10^{-4} \text{ mol} \]

\[ \text{molar mass} = \frac{\text{mass (in g)}}{\text{mol}} = \frac{0.0184 \text{ g}}{2.14 \times 10^{-4} \text{ mol}} = 86.0 \text{ g/mol} \]
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 C$_6$H$_{12}$.

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)**

\[
\text{CaO(s)} + \text{CO}_2(g) \rightarrow \text{CaCO}_3(s) \\
\text{BaO(s)} + \text{CO}_2(g) \rightarrow \text{BaCO}_3(s)
\]

**(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.

Initially:

\[
n_{\text{CO}_2} = \frac{PV}{RT} = \frac{746 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} \times (1.46 \text{ L})}{\frac{0.0821 \text{ L atm}}{\text{ mol} \cdot \text{K}}(35 + 273)\text{ K}} = 0.0567 \text{ mol CO}_2
\]

Remaining:

\[
n_{\text{CO}_2} = \frac{PV}{RT} = \frac{252 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} \times (1.46 \text{ L})}{\frac{0.0821 \text{ L atm}}{\text{ mol} \cdot \text{K}}(35 + 273)\text{ K}} = 0.0191 \text{ mol CO}_2
\]

Thus, the amount of CO$_2$ consumed in the reaction is: 

\(\text{0.0567 mol} - \text{0.0191 mol} = 0.0376 \text{ mol CO}_2\).

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:

\[
\text{mol CaO} + \text{mol BaO} = 0.0376 \text{ mol}
\]

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

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

\(0.01783x - 0.006523x + 0.0318 = 0.0376\)

\(x = 0.513 \text{ g} = \text{mass of CaO in the mixture}\)

mass of BaO in the mixture = \(4.88 - x = 4.37 \text{ g}\)

The percent compositions by mass in the mixture are:

\[
\text{CaO:} \frac{0.513 \text{ g}}{4.88 \text{ g}} \times 100\% = 10.5\% \\
\text{BaO:} \frac{4.37 \text{ g}}{4.88 \text{ g}} \times 100\% = 89.5\%
\]
5.126 (a) This is a Boyle’s law problem.

\[ P_{\text{tire}} V_{\text{tire}} = P_{\text{air}} V_{\text{air}} \]

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

\[ V_{\text{air}} = 4.90 \text{ L} \]

(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.

\[ P_{\text{pump}} V_{\text{pump}} = P_{\text{gauge}} V_{\text{gauge}} \]

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

\[ P_{\text{gauge}} = 0.33 \text{ atm} \]

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

5.127 (a) \[ \frac{188 \text{ g CO}}{1 \text{ h}} \times \frac{1 \text{ mol CO}}{28.01 \text{ g CO}} \times \frac{1 \text{ h}}{60 \text{ min}} = 0.112 \text{ mol CO / min} \]

(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 \(1.0 \times 10^{-3} \text{ atm} \) 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:

\[ P_{\text{CO}} = \frac{nRT}{V} = \frac{0.112 \text{ mol}}{5.3 \times 10^4 \text{ L}} \left( \frac{0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1}}{\text{mol} \cdot \text{K}} \right) \left( 20 + 273 \right) \text{K} \]

\[ = 5.1 \times 10^{-5} \text{ atm CO/min} \]

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

\[ ? \text{ min} = \frac{(1.0 \times 10^{-3} \text{ atm CO}) \times 1 \text{ min}}{5.1 \times 10^{-5} \text{ atm CO}} = 2.0 \times 10^4 \text{ min} \]

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.

\[ \frac{1 \text{ H atom}}{1 \text{ cm}^3} \times \frac{1 \text{ mol H}}{6.022 \times 10^{23} \text{ H atoms}} \times \frac{1 \text{ cm}^3}{1 \text{ mL}} \times \frac{1 \text{ mL}}{0.001 \text{ L}} = 2 \times 10^{-21} \text{ mol H} \]
The pressure of H is:

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

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

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

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

**5.129**

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

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

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.

\[
T = \frac{PV}{nR} = \left( \frac{210 \text{ mmHg} \times 1 \text{ atm}}{760 \text{ mmHg}} \right) \left( 1.00 \text{ L} \right) \left( 0.0147 \text{ mol} \right) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) = 229 \text{ K} = -44^\circ \text{C}
\]

(b) To determine the percent decrease in oxygen gas, let’s compare moles of \( \text{O}_2 \) at the top of Mt. Everest to the moles of \( \text{O}_2 \) at sea level.

\[
\frac{n_{\text{O}_2} \text{ (Mt. Everest)}}{n_{\text{O}_2} \text{ (sea level)}} = \frac{P_{\text{O}_2} \text{ (Mt. Everest)V}}{RT} = \frac{P_{\text{O}_2} \text{ (sea level)V}}{RT}
\]

\[
\frac{n_{\text{O}_2} \text{ (Mt. Everest)}}{n_{\text{O}_2} \text{ (sea level)}} = \frac{P_{\text{O}_2} \text{ (Mt. Everest)}}{P_{\text{O}_2} \text{ (sea level)}} = \frac{210 \text{ mmHg}}{760 \text{ mmHg}} = 0.276
\]

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 \( 3.9 \times 10^3 \text{ Pa} \) to units of mmHg:

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

\[
\text{Relative Humidity} = \frac{\text{partial pressure of water vapor fraction}}{\text{equilibrium vapor pressure}} \times 100\% = \frac{29 \text{ mmHg}}{31.82 \text{ mmHg}} \times 100\% = 91\%
\]
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:

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

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

\[ n = \frac{PV}{RT} = \frac{(1.0 \text{ atm})(5.2 \times 10^{-10} \text{ L})}{0.0821 \frac{\text{L atm}}{\text{mol K}} (37 + 273) \text{ K}} = 2.0 \times 10^{-11} \text{ mol of air} \]

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

\[ (2.0 \times 10^{-11} \text{ mol of air}) \times \frac{14\% \text{ oxygen}}{100\% \text{ air}} = 2.8 \times 10^{-12} \text{ mol O}_2 \]

Converting to O2 molecules:

\[ (2.8 \times 10^{-12} \text{ mol O}_2) \times \frac{6.022 \times 10^{23} \text{ O}_2 \text{ molecules}}{1 \text{ mol O}_2} = 1.7 \times 10^{12} \text{ O}_2 \text{ molecules} \]

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.

\[ V_{room} = (15.2 \text{ m})(6.6 \text{ m})(2.4 \text{ m}) \times \left( \frac{1 \text{ cm}}{0.01 \text{ m}} \right)^3 \times \frac{1 \text{ L}}{1000 \text{ cm}^3} = 2.4 \times 10^5 \text{ L} \]

\[ n_{Hg} = \frac{PV}{RT} = \frac{(1.7 \times 10^{-6} \text{ atm})(2.4 \times 10^5 \text{ L})}{0.0821 \frac{\text{L atm}}{\text{mol K}} (20 + 273) \text{ K}} = 0.017 \text{ mol Hg} \]

Converting to mass:

\[ ? \text{ g Hg} = 0.017 \text{ mol Hg} \times \frac{200.6 \text{ g Hg}}{1 \text{ mol Hg}} = 3.4 \text{ g Hg} \]

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

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

Yes, this far exceeds the air quality regulation of 0.050 mg Hg/m³ 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.

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

Some nitrogen oxides and their molar masses are:

\[
\begin{align*}
\text{NO} & \quad 30 \text{ g/mol} \\
\text{N}_2\text{O} & \quad 44 \text{ g/mol} \\
\text{NO}_2 & \quad 46 \text{ g/mol}
\end{align*}
\]

The nitrogen oxide is most likely NO\(_2\), although N\(_2\)O cannot be completely ruled out.

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

\[
\bar{\mathcal{M}} = \frac{\bar{P}V}{nRT} = \frac{(0.98 \text{ atm})(1.00 \text{ L})}{(0.0821 \frac{\text{L atm}}{\text{mol K}})(25 + 273) \text{K}} = 0.0401 \text{ mol}
\]

The average molar mass of the mixture is:

\[
\bar{\mathcal{M}} = \frac{2.7 \text{ g}}{0.0401 \text{ mol}} = 67 \text{ g/mol}
\]

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.

\[
\begin{align*}
X_{\text{NO}_2} \mathcal{M}_{\text{NO}_2} + X_{\text{N}_2\text{O}_4} \mathcal{M}_{\text{N}_2\text{O}_4} &= 67 \text{ g/mol} \\
X_{\text{NO}_2} (46.01 \text{ g/mol}) + (1 - X_{\text{NO}_2})(92.02 \text{ g/mol}) &= 67 \text{ g/mol} \\
46.01X_{\text{NO}_2} - 92.02X_{\text{NO}_2} + 92.02 &= 67 \\
X_{\text{NO}_2} &= 0.54 \quad \text{and} \quad X_{\text{N}_2\text{O}_4} = 1 - 0.54 = 0.46
\end{align*}
\]

Finally, the partial pressures are:

\[
\begin{align*}
P_{\text{NO}_2} &= X_{\text{NO}_2}P_T = (0.54)(0.98 \text{ atm}) = 0.53 \text{ atm} \\
P_{\text{N}_2\text{O}_4} &= X_{\text{N}_2\text{O}_4}P_T = (0.46)(0.98 \text{ atm}) = 0.45 \text{ atm}
\end{align*}
\]

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

\[
u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}} = \sqrt{\frac{3(8.314 \text{ J/mol-K})(1.7 \times 10^{-7} \text{ K})}{85.47 \times 10^{-3} \text{ kg/mol}}} = 7.0 \times 10^{-3} \text{ m/s}
\]

The mass of one Rb atom in kg is:

\[
\frac{85.47 \text{ g Rb}}{1 \text{ mol Rb}} \times \frac{1 \text{ mol Rb}}{6.022 \times 10^{23} \text{ Rb atoms}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 1.419 \times 10^{-25} \text{ kg/Rb atom}
\]

\[
\frac{1}{2} m \bar{u}^2 = \frac{1}{2} (1.419 \times 10^{-25} \text{ kg})(7.0 \times 10^{-3} \text{ m/s})^2 = 3.5 \times 10^{-30} \text{ J}
\]
5.137 First, calculate the moles of hydrogen gas needed to fill a 4.1 L life belt.

\[
\frac{PV}{RT} = \frac{(0.97 \text{ atm})(4.1 \text{ L})}{\left( \frac{0.0821 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right)(12 + 273)\text{K}} = 0.17 \text{ mol H}_2
\]

The balanced equation shows a mole ratio between \( \text{H}_2 \) and \( \text{LiH} \) of 1:1. Therefore, 0.17 mole of \( \text{LiH} \) is needed. Converting to mass in grams:

\[
? \text{ g LiH} = 0.17 \text{ mol LiH} \times \frac{7.949 \text{ g LiH}}{1 \text{ mol LiH}} = 1.4 \text{ g LiH}
\]

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

\[
V = \frac{nRT}{P} = \frac{(1 \text{ mol})(0.0821 \text{ L} \cdot \text{atm}/\text{mol} \cdot \text{K})(220 \text{ K})}{6.0 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}} = 2.3 \times 10^3 \text{ L}
\]

5.139 \( P_{\text{CO}_2} = (0.965) \times (9.0 \times 10^6 \text{ Pa}) = 8.7 \times 10^6 \text{ Pa} \)

\( P_{\text{N}_2} = (0.035) \times (9.0 \times 10^6 \text{ Pa}) = 3.2 \times 10^5 \text{ Pa} \)

\( P_{\text{SO}_2} = (1.5 \times 10^{-4}) \times (9.0 \times 10^6 \text{ Pa}) = 1.4 \times 10^3 \text{ Pa} \)

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 \times 28.02 g/mol) + (0.21 \times 32.00 g/mol) + (0.01 \times 39.95 g/mol) = 29 g/mol

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

Now, we can calculate the volume of the bulb.

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

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. \[ u_{\text{rms}} = \sqrt{\frac{(3)(8.314 \text{ J/mol K})(74 + 273)K}{4.003 \times 10^{-2} \text{ kg/mol}}} = 1.47 \times 10^3 \text{ m/s} \]

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

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

(b) First, calculate the moles of air exhaled in every breath. (500 mL = 0.500 L)
\[ n = \frac{PV}{RT} = \frac{(1 \text{ atm})(0.500 \text{ L})}{(0.0821 \text{ L atm/mol K})(37 + 273)\text{K}} = 1.96 \times 10^{-2} \text{ mol air/breath} \]
Next, convert to molecules of air per breath.
\[ 1.96 \times 10^{-2} \text{ mol air/breath} \times \frac{6.022 \times 10^{23} \text{ molecules air}}{1 \text{ mol air}} = 1.18 \times 10^{22} \text{ molecules/breath} \]

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

(d) Fraction of molecules in the atmosphere exhaled by Mozart is:
\[ \frac{2.60 \times 10^{30} \text{ molecules}}{1.09 \times 10^{44} \text{ molecules}} = 2.39 \times 10^{-14} \]
Or,
\[ \frac{1}{2.39 \times 10^{-14}} = 4.18 \times 10^{13} \]
Thus, about 1 molecule of air in every $4 \times 10^{13}$ molecules was exhaled by Mozart.

In a single breath containing $1.18 \times 10^{22}$ molecules, we would breathe in on average:
\[ (1.18 \times 10^{22} \text{ molecules}) \times \frac{1 \text{ Mozart air molecule}}{4 \times 10^{13} \text{ air molecules}} = 3 \times 10^8 \text{ molecules that Mozart exhaled} \]

(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\textsubscript{2} at 25°C.
\[ u_{\text{rms}}(\text{N}_2) = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3(8.314 \text{ J/mol K})(298 \text{ K})}{0.02802 \text{ kg/mol}}} = 515 \text{ m/s} \]
Now, we can calculate the temperature at which He atoms will have this same root-mean-square speed.

\[ u_{\text{rms}}(\text{He}) = \sqrt{\frac{3RT}{M}} \]

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

\[ T = 42.6 \text{ K} \]

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

\[ \frac{n}{V} = \frac{P}{RT} = \frac{1.0 \text{ atm}}{(0.0821 \text{ L·atm/mol·K})(100 + 273)K} = 0.033 \text{ mol/L} \]

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 (m³).

\[ \frac{0.033 \text{ mol}}{1 \text{ L}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \times \frac{1000 \text{ L}}{1 \text{ m}^3} = 2.0 \times 10^{25} \text{ molecules/m}^3 \]

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:

\[ \left( \frac{2.0 \times 10^{25} \text{ molecules}}{1 \text{ m}^3} \right)^{\frac{1}{3}} = 2.7 \times 10^8 \frac{\text{molecules}}{\text{m}} \]

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

\[ \frac{1 \text{ m}}{2.70 \times 10^8} = 3.7 \times 10^{-9} \text{ m} = 3.7 \text{ nm} \]

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.

\[ \frac{0.96 \text{ g/cm}^3}{1 \text{ cm}^3} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol H}_2\text{O}} \times \left( \frac{100 \text{ cm}^3}{1 \text{ m}^3} \right)^3 = 3.2 \times 10^{28} \text{ molecules/m}^3 \]

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.

\[ \left( \frac{3.2 \times 10^{28} \text{ molecules}}{1 \text{ m}^3} \right)^{\frac{1}{3}} = 3.2 \times 10^9 \frac{\text{molecules}}{\text{m}} \]

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

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

\[
29 \text{ g/mol} = 0.029 \text{ kg/mol}
\]

\[
5.0 \text{ km} = 5.0 \times 10^3 \text{ m}
\]

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

\[
P = P_0 e^{-\frac{gh}{RT}}
\]

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

\[
P = 0.54 \text{ atm}
\]

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.

\[
P = \frac{nRT}{V} = \frac{(2.14 \text{ mol})(0.0821 \text{ L atm/mol K})(182 + 273)\text{K}}{8.00 \text{ L}} = 9.99 \text{ atm}
\]
The relative rates of effusion of the two gases will allow the calculation of the mole fraction of each gas as it passes through the orifice. Using Graham’s law of effusion,

$$\frac{n_{H_2}}{n_{D_2}} = \sqrt{\frac{M_{D_2}}{M_{H_2}}}$$

$$\frac{n_{H_2}}{n_{D_2}} = \sqrt{\frac{4.028 \text{ g/mol}}{2.016 \text{ g/mol}}}$$

$$\frac{n_{H_2}}{n_{D_2}} = 1.414$$

This calculation shows that H$_2$ will effuse 1.414 times faster than D$_2$. If, over a given amount of time, 1 mole of D$_2$ effuses through the orifice, 1.414 moles of H$_2$ will effuse during the same amount of time. We can now calculate the mole fraction of each gas.

$$X_{H_2} = \frac{\text{mol H}_2}{\text{total mol}} = \frac{1.414 \text{ mol}}{(1 + 1.414) \text{ mol}} = 0.5857$$

Because this is a two-component mixture, the mole fraction of D$_2$ is:

$$X_{D_2} = 1 - 0.5857 = 0.4143$$

CaCO$_3$(s) + 2HCl(aq) $\longrightarrow$ CaCl$_2$(aq) + CO$_2$(g) + H$_2$O(l)

MgCO$_3$(s) + 2HCl(aq) $\longrightarrow$ MgCl$_2$(aq) + CO$_2$(g) + H$_2$O(l)

First, we need to find the number of moles of CO$_2$ produced in the reaction. We can do this by using the ideal gas equation. We carry an extra significant figure throughout this calculation to minimize rounding errors.

$$n_{CO_2} = \frac{PV}{RT} = \frac{(1.12 \text{ atm})(1.73 \text{ L})}{(0.0821 \text{ L atm/mol} \cdot \text{K})(48 + 273)\text{K}} = 0.07352 \text{ mol CO}_2$$

Since the mole ratio between CO$_2$ and both reactants (CaCO$_3$ and MgCO$_3$) is 1:1, 0.07352 mole of the mixture must have reacted. We can write:

$$\text{mol CaCO}_3 + \text{mol MgCO}_3 = 0.07352 \text{ mol}$$

Let $x$ g = mass of CaCO$_3$ in the mixture, then $(6.26 - x)$ g = mass of MgCO$_3$ in the mixture. We can write:

$$\left[ x \text{ g CaCO}_3 \times \frac{1 \text{ mol CaCO}_3}{100.1 \text{ g CaCO}_3} \right] + \left[ (6.26 - x) \text{ g MgCO}_3 \times \frac{1 \text{ mol MgCO}_3}{84.32 \text{ g MgCO}_3} \right] = 0.07352 \text{ mol}$$

$$0.009990x - 0.01186x + 0.07424 = 0.07352$$

$$x = 0.385 \text{ g} = \text{mass of CaCO}_3 \text{ in the mixture}$$

mass of MgCO$_3$ in the mixture = $6.26 - x = 5.88 \text{ g}$
The percent compositions by mass in the mixture are:

\[
\text{CaCO}_3: \frac{0.385 \text{ g}}{6.26 \text{ g}} \times 100\% = 6.15\%
\]

\[
\text{MgCO}_3: \frac{5.88 \text{ g}}{6.26 \text{ g}} \times 100\% = 93.9\%
\]

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

From the amount of \( \text{H}_2(\text{g}) \) 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.

\[
\begin{align*}
\text{n}_{\text{H}_2} &= \frac{PV_{\text{H}_2}}{RT} \\
\text{n}_{\text{H}_2} &= \frac{728 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} \times (1.26 \text{ L})}{0.0821 \frac{\text{L atm}}{\text{mol K}}} (22 + 273)\text{K} \\
&= 0.0498 \text{ mol H}_2
\end{align*}
\]

Since the mole ratio between \( \text{H}_2 \) 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:

\[
\text{mass % Zn} = \frac{\text{mass Zn}}{\text{mass sample}} \times 100\% = \frac{3.26 \text{ g}}{6.11 \text{ g}} \times 100\% = 53.4\%
\]

5.151 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:

\[
P_T = P_{\text{air}} + P_{\text{acetone}}
\]

\[
P_{\text{air}} = P_T - P_{\text{acetone}} = 750 \text{ mmHg} - 400 \text{ mmHg} = 350 \text{ mmHg}
\]

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:

\[ P_T = P_{\text{air}} + P_{\text{acetone}} \]

\[ P_T = 672 \text{ mmHg} + 400 \text{ mmHg} = 1072 \text{ mmHg} \]

5.152 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 CO₂. 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.

\[ \frac{r_{\text{He}}}{r_{\text{mix}}} = \sqrt{\frac{M_{\text{mix}}}{M_{\text{He}}}} \]

\[ M_{\text{mix}} = \left( \frac{r_{\text{He}}}{r_{\text{mix}}} \right)^2 M_{\text{He}} \]

\[ M_{\text{mix}} = \left( \frac{29.7 \text{ mL}}{2.00 \text{ min}} \times \frac{4.003 \text{ g/mol}}{10.0 \text{ mL} / \text{2.00 min}} \right)^2 = 35.31 \text{ g/mol} \]

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.

\[ X_{\text{CO}}M_{\text{CO}} + X_{\text{CO}_2}M_{\text{CO}_2} = M_{\text{mix}} \]

\[ X_{\text{CO}}M_{\text{CO}} + (1 - X_{\text{CO}})M_{\text{CO}_2} = M_{\text{mix}} \]

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

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

\[ 16.00X_{\text{CO}} = 8.70 \]

\[ X_{\text{CO}} = 0.544 \]

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

\[ \% \text{ of CO by volume} = 54.4\% \]

\[ \% \text{ of CO}_2 \text{ by volume} = 1 - \% \text{ of CO by volume} = 45.6\% \]

5.153 (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.

5.154 The reactions are:

$$\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$$

$$2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}$$

For a given volume and temperature, $n \propto P$. 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$_2$.

$$\left( \frac{R_{\text{CH}_4}}{1 \text{ mol CO}_2} \times \frac{1 \text{ mol CH}_4}{1 \text{ mol CO}_2} \right) + \left( \frac{R_{\text{C}_2\text{H}_6}}{4 \text{ mol CO}_2} \times \frac{2 \text{ mol C}_2\text{H}_6}{4 \text{ mol CO}_2} \right) = 356 \text{ mmHg CO}_2$$

(1) $P_{\text{CH}_4} + 2P_{\text{C}_2\text{H}_6} = 356 \text{ mmHg}$

Also,

(2) $P_{\text{CH}_4} + P_{\text{C}_2\text{H}_6} = 294 \text{ mmHg}$

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

$$P_{\text{CH}_4} = 356 - 294 = 62 \text{ mmHg}$$

$$P_{\text{C}_2\text{H}_6} = 294 - 62 = 232 \text{ mmHg}$$

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

$$X_{\text{CH}_4} = \frac{232}{294} = 0.789 \quad X_{\text{C}_2\text{H}_6} = \frac{62}{294} = 0.211$$

5.155 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.

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

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

This is eight times the volume of an individual molecule.
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, \( N_A \).

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

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.

**5.157** 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:**

\[
P = \frac{nRT}{V} = \frac{(5.00 \text{ mol})(0.0821 \text{ L atm/mol K})(300 \text{ K})}{1.92 L} = 64.1 \text{ atm}
\]

**van der Waals:**

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

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

\[
P = 71.0 \text{ atm} - 28.3 \text{ atm} = 42.7 \text{ atm}
\]

The \( a \) and \( b \) values used are from Table 5.4 of the text.

The percent error in using the ideal gas equation is:

\[
\% \text{ error} = \left| \frac{P_{\text{vdw}} - P_{\text{ideal}}}{P_{\text{vdw}}} \right| \times 100%
\]

\[
\% \text{ error} = \left| \frac{42.7 \text{ atm} - 64.1 \text{ atm}}{42.7 \text{ atm}} \right| \times 100% = 50.1\%
\]

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

\[
u_{\text{rms}} = \sqrt{\frac{3RT}{M}}
\]

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

5.159 (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·m²/s².

$$u_{mp} = \sqrt{\frac{2RT}{M}}$$
$$u_{rms} = \sqrt[3]{\frac{3RT}{M}}$$

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

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

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.

$$\frac{u_{mp}}{u_{rms}} = \sqrt[3]{\frac{2RT}{3RT}} = \sqrt[3]{\frac{2}{3}} = 0.816$$

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 $T_2$.

$$\frac{u_{mp}(1)}{u_{mp}(2)} = \sqrt{\frac{2RT_1}{M}} = \sqrt[3]{\frac{2RT_1}{3RT_2}}$$
$$\frac{u_{mp}(1)}{u_{mp}(2)} = \sqrt[3]{\frac{T_1}{T_2}}$$

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 $T_2$.

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

$$0.5 = \sqrt{\frac{300}{T_2}}$$

$$T_2 = 1200 \text{ K}$$
5.160 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.

5.161 First, the empirical formula of the gaseous hydrocarbon can be determined from the amount of CO₂ and H₂O produced. We assume that all the carbon in the hydrocarbon ends up in CO₂ and all the hydrogen ends up in water.

\[
\begin{align*}
205.1 \text{ g CO}_2 & \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 4.660 \text{ mol C} \\
168.0 \text{ g H}_2\text{O} & \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 18.65 \text{ mol H}
\end{align*}
\]
This gives a correct formula for the compound, C₄H₈O. However, it is not an empirical formula. We divide each subscript by the smallest number moles to arrive at the empirical formula, CH₄.

Next, from the pressure, volume, and temperature, we can determine the moles of gaseous hydrocarbon in the container prior to combustion.
\[
n = \frac{PV}{RT} = \frac{(6.63 \text{ atm})(20.2 \text{ L})}{(0.0821 \text{ L atm/mol K})(350 \text{ K})} = 4.66 \text{ mol}
\]
In the reaction, we start with 4.66 moles of the hydrocarbon and 4.660 moles of CO₂ are produced.
\[
\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}
\]
This indicates that for every mole of the hydrocarbon reacted, 1 mole of CO₂ is produced. Or, we can also say that the mole ratio between CH₄ and CO₂ is 1:1. The molecular formula of the hydrocarbon is CH₄.

5.162 (i) Bulb (b) contains the same number of particles as (a), but is half the volume. The pressure will be double that of A.
\[
P_B = (2)(4.0 \text{ atm}) = 8.0 \text{ atm}
\]
The volume of bulb (c) is the same as bulb (a), but there are 12 particles in (c) while there are 9 particles in (a). The pressure is directly proportional to the number of moles of gas (or particles of gas) at the same temperature and volume.

\[ P_c = \left( \frac{12}{9} \right) (4.0 \text{ atm}) = 5.3 \text{ atm} \]

(ii) When the valves are opened at constant temperature, the gases expand to fill the entire container. For example, in bulb (a), the pressure before opening the valve is 4.0 atm and the volume is 4.0 L. After the valves are opened, the volume is now 10 L (4 L + 2 L + 4 L). We use Boyle’s law to calculate the partial pressure of each sample of gas after the valves are opened.

(a) \[ P_1 V_1 = P_2 V_2 \]
\[ (4.0 \text{ atm})(4.0 \text{ L}) = P_2(10.0 \text{ L}) \]
\[ P_2 = 1.6 \text{ atm} \]

(b) \[ P_1 V_1 = P_2 V_2 \]
\[ (8.0 \text{ atm})(2.0 \text{ L}) = P_2(10.0 \text{ L}) \]
\[ P_2 = 1.6 \text{ atm} \]

(c) \[ P_1 V_1 = P_2 V_2 \]
\[ (5.3 \text{ atm})(4.0 \text{ L}) = P_2(10.0 \text{ L}) \]
\[ P_2 = 2.1 \text{ atm} \]

The total pressure in the container is the sum of the partial pressures (Dalton’s Law).

\[ P_{\text{total}} = 1.6 \text{ atm} + 1.6 \text{ atm} + 2.1 \text{ atm} = 5.3 \text{ atm} \]

There are 15 particles of gas A and 15 particles of gas B in the container. Therefore, the partial pressure of each gas will be half the total pressure of 5.3 atm.

\[ P_A = P_B = 2.65 \text{ atm} \]

5.163 (a) \[ 2\text{KO}_2 + 2\text{H}_2\text{O} \rightarrow \text{O}_2 + 2\text{KOH} + \text{H}_2\text{O}_2 \]

(b) First, we calculate the density of O\(_2\) in KO\(_2\) using the mass percentage of O\(_2\) in the compound.
\[
\frac{32.00 \text{ g O}_2}{71.10 \text{ g KO}_2} \times \frac{2.15 \text{ g}}{1 \text{ cm}^3} = 0.968 \text{ g/cm}^3
\]

Now, we can use Equation (5.11) of the text to calculate the pressure of oxygen gas that would have the same density as that provided by KO\(_2\).

\[
P = \frac{dRT}{M} = \left( \frac{968 \text{ g}}{1 \text{ L}} \right) \left( \frac{0.0821 \text{ L atm/mol} \cdot \text{K}}{1 \text{ mol} \cdot \text{K}} \right) = 727 \text{ atm}
\]

Obviously, using O\(_2\) instead of KO\(_2\) is not practical.
### ANSWERS TO REVIEW OF CONCEPTS

**Section 5.2 (p. 179)**  
It would be easier to drink water with a straw at the foot of Mt. Everest because the atmospheric pressure is greater there, which helps to push the water up the straw.

**Section 5.4 (p. 187)**  
(b)

**Section 5.6 (p. 201)**  
Blue sphere: **0.43 atm**.  Green sphere: **1.3 atm**.  Red sphere: **0.87 atm**.

**Section 5.7 (p. 211)**  
(a) **No**.  Collisions among He atoms and between He atoms and the walls only result in the transfer of energy.  Therefore, the total energy of the He atoms remains unchanged as long as the temperature is kept constant.

(b) UF₆ is heavier but moves more slowly than H₂.  Thus, according to Equation (5.15) of the text, the average kinetic energy of these two gases is the same (assuming ideal behavior).