Answers and Solutions to Text Problems

8.1 a. Gaseous particles have greater kinetic energies at higher temperatures. Because kinetic energy is a measure of the energy of motion, the gaseous particles must be moving faster at higher temperatures than at lower values.
b. Because particles in a gas are very far apart, gases can be easily compressed without the particles bumping into neighboring gas particles. Neighboring particles are much closer together in solids and liquids, and they will “bump” into each other and repel each other if the sample is compressed.
c. Because the particles of a gas are very far apart, only a small amount of mass (due to the gas particles themselves) is found in a given volume of space.

8.2 a. At the higher temperature in the fire, the number of collisions against the container’s walls increase because the gaseous particles have greater velocities. This increases the pressure in the can, and if that pressure exceeds what the container can endure, then the container will explode.
b. The particles of a gas move faster at higher temperatures. This causes the particles to spread farther apart, reducing the gas sample’s density. The density of air in the balloon is less, which causes it to rise until its density becomes equal to the surrounding atmosphere.
c. Because there are fewer gaseous particles in a sample of air at a higher altitude, there will be fewer collisions and thus a lower pressure.

8.3 a. temperature b. volume c. amount of gas d. pressure

8.4 a. temperature b. pressure c. volume d. amount of gas

8.5 Some units used to describe the pressure of a gas are pounds per square inch (lb/in.² which is also abbreviated as psi), atmospheres (abbreviated atm), torr, mm Hg, in. Hg, and pascals.

8.6 Statements a, d, and e describe the pressure of a gas.

8.7 a. 2.00 atm \times \frac{760 \text{ torr}}{1 \text{ atm}} = 1520 \text{ torr}
b. 2.00 atm \times \frac{14.7 \text{ lb/in.}²}{1 \text{ atm}} = 29.4 \text{ lb/in.}²
c. 2.00 atm \times \frac{760 \text{ mm Hg}}{1 \text{ atm}} = 1520 \text{ mm Hg}

8.8 a. 467 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 0.614 \text{ atm}
b. 467 \text{ mm Hg} \times \frac{1 \text{ torr}}{1 \text{ mm Hg}} = 467 \text{ torr}
c. 467 \text{ mm Hg} \times \frac{1 \text{ cm}}{10 \text{ mm}} \times \frac{1 \text{ in. Hg}}{2.54 \text{ cm}} = 18.4 \text{ in. Hg}

8.9 The gases in the diver’s lungs (and dissolved in the blood) will expand because pressure decreases as the diver ascends. Unless the diver exhales, the expanding gases could rupture the membranes in the lung tissues. In addition, the formation of gas bubbles in the bloodstream could cause “the bends.”
Chapter 8 Answers and Solutions

8.10 According to Boyle’s law, gases expand as the pressure is decreased. Because atmospheric pressure decreases as altitude increases, the volume of gas sealed in the bag of chips will increase at higher altitudes.

8.11 During expiration the volume (capacity) of the lungs is decreased.

8.12 A respirator inflates the lungs with air that is rich in oxygen, (and has a low concentration of carbon dioxide). This allows the blood to rid itself of the waste CO₂ from the body’s cells and pick up more O₂ to be delivered to the body’s cells as the blood circulates.

8.13 a. According to Boyle’s law, for the pressure to increase while temperature and quantity of gas remains constant, the gas volume must decrease. Thus, cylinder A would represent the final volume.

<table>
<thead>
<tr>
<th></th>
<th>Initial</th>
<th>Final</th>
</tr>
</thead>
<tbody>
<tr>
<td>P</td>
<td>650 mm Hg</td>
<td>1.2 atm</td>
</tr>
<tr>
<td>V</td>
<td>220 mL</td>
<td>160 mL</td>
</tr>
</tbody>
</table>

Because \( P_1V_2 = P_2V_1 \), then \( V_2 = \frac{P_1V_1}{P_2} \)

\[
V_2 = 220 \text{ mL} \times \frac{650 \text{ mm Hg}}{1.2 \text{ atm}} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 160 \text{ mL}
\]

8.14 a. According to Boyle’s law, the volume must increase when the pressure is decreased, so diagram C should represent the balloon’s final volume at an increased altitude.

b. When there is no change in the pressure, there is no change in the volume. Diagram B would represent the “final” volume of the balloon.

c. With an increase in pressure there will be a decrease in volume. Thus diagram A would describe the balloon’s final volume.

8.15 a. The pressure doubles when the volume is halved.

b. The pressure falls to one-third the initial pressure when the volume expands to three times its initial volume.

c. The pressure increases to ten times the original pressure when the volume decreases to 1/10 of its initial volume.

8.16 a. The volume decreases to one-third the initial volume, when pressure increases to three times its initial pressure.

b. The volume doubles when pressure falls to one-half its initial pressure.

c. The volume is 5 times greater when pressure is decreased to 1/5 its initial pressure.

8.17 From Boyle’s law we know that pressure is inversely related to volume. (For example, the pressure increases when the volume decreases.)

a. Volume increases; pressure must decrease.

\[
655 \text{ mm Hg} \times \frac{10.0 \text{ L}}{20.0 \text{ L}} = 328 \text{ mm Hg}
\]

b. Volume decreases; pressure must increase.

\[
655 \text{ mm Hg} \times \frac{10.0 \text{ L}}{2.50 \text{ L}} = 2620 \text{ mm Hg}
\]

c. The mL units must be converted to L for unit cancellation in the calculation, and because the volume decreases; pressure must increase.

\[
655 \text{ mm Hg} \times \frac{10.0 \text{ L}}{1500 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 4400 \text{ mm Hg}
\]
d. The mL units must be converted to L for unit cancellation in the calculation, and because the volume decreases; pressure must increase.

\[ 655 \text{ mm Hg} \times \frac{10.0 \text{ L}}{120. \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 55000 \text{ mm Hg} \]

8.18 From Boyle's Law we know that pressure is inversely related to volume.

a. \( 1.20 \text{ atm} \times \frac{5.00 \text{ L}}{1.00 \text{ L}} = 6.00 \text{ atm} \)

b. The mL units must be converted to L for unit cancellation in the calculation.

\( 1.20 \text{ atm} \times \frac{5.00 \text{ L}}{2500 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 2.4 \text{ atm} \)

\( 1.20 \text{ atm} \times \frac{5.00 \text{ L}}{750 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 8.0 \text{ atm} \)

\( 1.20 \text{ atm} \times \frac{5.00 \text{ L}}{8.00 \text{ L}} = 0.750 \text{ atm} \)

8.19 \( 760 \text{ mm Hg} \times \frac{4.5 \text{ L}}{2.0 \text{ L}} = 1700 \text{ mm Hg} \)

8.20 \( 15.0 \text{ atm} \times \frac{20.0 \text{ L}}{300.0 \text{ L}} = 1.00 \text{ atm} \)

8.21 From Boyle's Law we know that pressure is inversely related to volume.

a. Pressure increases; volume must decrease.

\( 50.0 \text{ L} \times \frac{760 \text{ mm Hg}}{1500 \text{ mm Hg}} = 25 \text{ L} \)

b. The mm Hg units must be converted to atm for unit cancellation in the calculation.

\( 50.0 \text{ L} \times \frac{760 \text{ mm Hg}}{2.0 \text{ atm}} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 25 \text{ L} \)

c. The mm Hg units must be converted to atm for unit cancellation in the calculation.

\( 50.0 \text{ L} \times \frac{760 \text{ mm Hg}}{0.500 \text{ atm}} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 100. \text{ L} \)

d. The mm Hg units must be converted to torr for unit cancellation in the calculation.

\( 50.0 \text{ L} \times \frac{760 \text{ mm Hg}}{850 \text{ torr}} \times \frac{1 \text{ torr}}{1 \text{ mm Hg}} = 45 \text{ L} \)

8.22 From Boyle's Law we know that pressure is inversely related to volume.

a. \( 25 \text{ mL} \times \frac{0.80 \text{ atm}}{0.40 \text{ atm}} = 50. \text{ mL} \)

b. \( 25 \text{ mL} \times \frac{0.80 \text{ atm}}{2.00 \text{ atm}} = 10. \text{ mL} \)

c. \( 25 \text{ mL} \times \frac{0.80 \text{ atm}}{2500 \text{ mm Hg}} \times \frac{760 \text{ mm Hg}}{1 \text{ atm}} = 6.1 \text{ mL} \)

d. \( 25 \text{ mL} \times \frac{0.80 \text{ atm}}{80.0 \text{ torr}} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 190 \text{ mL} \)

8.23 \( 5.0 \text{ L} \times \frac{5.0 \text{ atm}}{1.0 \text{ atm}} = 25 \text{ L} \)

8.24 \( 615 \text{ mL} \times \frac{760 \text{ mm Hg}}{752 \text{ mm Hg}} = 620 \text{ mL} \)
8.25  
a. Inspiration begins when the diaphragm flattens causing the lungs to expand. The increased volume reduces the pressure in the lungs such that air flows into the lungs.

b. Expiration occurs as the diaphragm relaxes causing a decrease in the volume of the lungs. The pressure of the air in the lungs increases and air flows out of the lungs.

c. Inspiration occurs when the pressure in the lungs is less than the pressure of the air in the atmosphere.

8.26  
a. Expiration occurs when the diaphragm relaxes and decreases the volume of the lungs.

b. Inspiration occurs when the lungs expand.

c. Expiration occurs when the pressure in the lungs is higher than the pressure in the atmosphere.

8.27  
According to Charles' law, there is a direct relationship between temperature and volume. For example, volume increases when temperature increases while the pressure and amount of gas remains constant.

a. Diagram C describes an increased volume corresponding to an increased temperature.

b. Diagram A describes a decreased volume corresponding to a decrease in temperature.

c. Diagram B shows no change in volume, which corresponds to no change in temperature.

8.28  
According to Charles’ law, an increase in temperature will give an increase in volume.

a. Because the gas warms, its volume must also increase.

b. Because the warm gas will cool, its volume must decrease.

C. Because the gas warms, its volume must increase.

8.29  
Heating a gas in a hot air balloon increases the volume of gas, which reduces its density and allows the balloon to rise above the ground.

8.30  
As the temperature decreases, the volume of the gas in the tire decreases, which makes the tire appear lower or flat in the morning.

8.31  
According to Charles’ law, gas volume is directly proportional Kelvin temperature when P and n are constant. In all gas law computations, temperatures must be in Kelvin units. (Temperatures in °C are converted to K by the addition of 273.)

a. When temperature decreases, volume must also decrease.

\[
75°C + 273 = 348 K, 55°C + 273 = 328 K \\
2500 mL × \frac{328 K}{348 K} = 2400 mL
\]

b. When temperature increases, volume must also increase.

\[
2500 mL × \frac{680 K}{348 K} = 4900 mL
\]

c. 
\[
-25°C + 273 = 248 K \\
2500 mL × \frac{248 K}{348 K} = 1800 mL
\]

d. 
\[
2500 mL × \frac{240 K}{348 K} = 1700 mL
\]

8.32  
According to Charles’ law, a change in a gas’s volume is directly proportional to the change in its Kelvin temperature. In all gas law computations, temperatures must be in Kelvin units. (Temperatures in °C are converted to K by the addition of 273.)

a. 0°C + 273 = 273 K

\[
273 K × \frac{10.0 L}{4.00 L} = 683 K, 683 K - 273 = 410°C
\]
According to Gay-Lussac’s law, the pressure of a gas is directly related to the temperature when volume and the number of moles remain constant.

According to Gay-Lussac’s law, temperature is directly related to pressure. For example, temperature increases when the pressure increases. In all gas law computations, temperatures must be in Kelvin units. (Temperatures in °C are converted to K by the addition of 273.)

**a.** 155°C + 273 = 428 K

\[
1200 \text{ torr} \times \frac{273 \text{ K}}{428 \text{ K}} = 770 \text{ torr}
\]

**b.** 12°C + 273 = 285 K

\[
1.40 \text{ atm} \times \frac{308 \text{ K}}{285 \text{ K}} = 1.51 \text{ atm}
\]

According to Gay-Lussac’s law, pressure is directly related to temperature. For example, pressure increases when the temperature increases. In all gas law computations, temperatures must be in Kelvin units. (Temperatures in °C are converted to K by the addition of 273.)

**a.** 0°C + 273 = 273 K

\[
273 \text{ K} \times \frac{1500 \text{ torr}}{250 \text{ torr}} = 1600 \text{ K}
\]

\[
1600 \text{ K} - 273 = 1300°C
\]

**b.** (This is not a Gay-Lussac law problem! It is a Charles law problem, but is solved in an analogous fashion to a Gay-Lussac problem.)

\[
-10.0°C + 273 = 263 \text{ K} \quad 37°C + 273 = 310. \text{ K}
\]

\[
500.0 \text{ mL} \times \frac{310. \text{ K}}{263 \text{ K}} = 589 \text{ mL}
\]

Boiling would occur in b and c because in both of these the vapor pressure equals the atmospheric (external) pressure.

Water boils at temperatures less than 100°C because the atmospheric pressure is less than one atmosphere on a mountain. The boiling point is the temperature at which the vapor pressure of a liquid becomes equal to the external (in this case, atmospheric) pressure.
b. The pressure inside a pressure cooker is greater than one atmosphere; therefore water boils above 100°C. Foods cook faster at higher temperatures.

8.42 a. The wind carries away the water vapor molecules, reducing the humidity near the damp towel. 
b. At an external pressure of 2.0 atm water will boil at 120°C.

8.43 Boyle’s, Charles’, and Gay-Lussac’s laws are united into the combined gas law. 
\[ \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \]

8.44 

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

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

8.45 \[ T_1 = 25°C + 273 = 298 \text{ K}; \ V_1 = 6.50 \text{ L}; \ P_1 = 845 \text{ mm Hg (1.11 atm)} \]

a. \[ T_2 = 325 \text{ K}; \ V_2 = 1.85 \text{ L} \]
\[ 1.11 \text{ atm} \times \frac{6.50 \text{ L}}{1.85 \text{ L}} \times \frac{325 \text{ K}}{298 \text{ K}} = 4.25 \text{ atm} \]

b. \[ T_2 = 12°C + 273 = 285 \text{ K}; \ V_2 = 2.25 \text{ L} \]
\[ 1.11 \text{ atm} \times \frac{6.50 \text{ L}}{2.25 \text{ L}} \times \frac{285 \text{ K}}{298 \text{ K}} = 3.07 \text{ atm} \]

c. \[ T_2 = 47°C + 273 = 320 \text{ K}; \ V_2 = 12.8 \text{ L} \]
\[ 1.11 \text{ atm} \times \frac{6.50 \text{ L}}{12.8 \text{ L}} \times \frac{320 \text{ K}}{298 \text{ K}} = 0.605 \text{ atm} \]

8.46 \[ T_1 = 112°C + 273 = 385 \text{ K}; \ P_1 = 1.20 \text{ atm}; \ V_1 = 735 \text{ mL} \]

a. \[ T_2 = 281 \text{ K}; \ P_2 = 658 \text{ mm Hg (0.866 atm)} \ V_2 = \text{? mL} \]
\[ V_2 = 735 \text{ mL} \times \frac{1.20 \text{ atm}}{0.866 \text{ atm}} \times \frac{281 \text{ K}}{385 \text{ K}} = 743 \text{ mL} \]

b. \[ T_2 = 75°C + 273 = 348 \text{ K}; \ P_2 = 0.55 \text{ atm} \ V_2 = \text{? mL} \]
\[ V_2 = 735 \text{ mL} \times \frac{1.20 \text{ atm}}{0.55 \text{ atm}} \times \frac{348 \text{ K}}{385 \text{ K}} = 1450 \text{ mL} \]

c. \[ T_2 = -15°C + 273 = 258 \text{ K}; \ P_2 = 15.4 \text{ atm} \ V_2 = \text{? mL} \]
\[ V_2 = 735 \text{ mL} \times \frac{1.20 \text{ atm}}{15.4 \text{ atm}} \times \frac{258 \text{ K}}{385 \text{ K}} = 3.84 \text{ mL} \]

8.48 \[ T_1 = 8°C + 273 = 281 \text{ K}; \ V_1 = 50.0 \text{ mL}; \ P_1 = 3.00 \text{ atm} \]
\[ T_2 = 37°C + 273 = 310. \text{ K}; \ V_2 = 150.0 \text{ mL} \ P_2 = \text{?} \]
\[ P_2 = 3.00 \text{ atm} \times \frac{50.0 \text{ mL}}{150.0 \text{ mL}} \times \frac{310 \text{ K}}{281 \text{ K}} = 1.10 \text{ atm} \]

8.49 Addition of more air molecules to a tire or basketball will increase its volume.

8.50 Air molecules are escaping from the balloon, (it is deflating), as it flies around the room.
According to Avogadro’s law, a change in a gas’s volume is directly proportional to the change in the number of moles of gas.

a. 8.00 L × \( \frac{2.00 \text{ mole}}{4.00 \text{ mole}} \) = 4.00 L

b. 25.0 g neon × \( \frac{1 \text{ mole neon}}{20.2 \text{ g neon}} \) = 1.24 mole Ne added

\[
1.50 \text{ mole} + 1.24 \text{ mole} = 2.74 \text{ mole}
\]

8.00 L × \( \frac{2.74 \text{ mole}}{1.50 \text{ mole}} \) = 14.6 L

c. 1.50 mole + 3.50 mole = 5.00 mole

8.00 L × \( \frac{5.00 \text{ mole}}{1.50 \text{ mole}} \) = 26.7 L

8.52

a. 4.80 g O\(_2\) × \( \frac{1 \text{ mole O}_2}{32.0 \text{ g O}_2} \) = 0.150 mole O\(_2\) New moles = 0.650 mole O\(_2\)

15.0 L × \( \frac{0.650 \text{ mole O}_2}{0.150 \text{ mole O}_2} \) = 65.0 L

b. 0.150 mole × \( \frac{10.0 \text{ L}}{15.0 \text{ L}} \) = 0.100 mole O\(_2\) remain;

0.150 mole O\(_2\) - 0.100 mole O\(_2\) = 0.050 mole O\(_2\) removed

c. 4.00 g He × \( \frac{1 \text{ mole He}}{4.00 \text{ g He}} \) = 1.00 mole He New moles = 1.15 moles of gases

15.0 L × \( \frac{1.15 \text{ moles gas}}{0.15 \text{ mole O}_2} \) = 115 L

8.53 At STP, the molar volume of any gas is 22.4 L per mole.

a. 44.8 L × \( \frac{1 \text{ mole O}_2}{22.4 \text{ L}} \) = 2.00 mole O\(_2\)

b. 4.00 L × \( \frac{1 \text{ mole CO}_2}{22.4 \text{ L}} \) = 0.179 mole CO\(_2\)

c. 6.40 g O\(_2\) × \( \frac{1 \text{ mole O}_2}{32.0 \text{ g O}_2} \) × \( \frac{22.4 \text{ L}}{1 \text{ mole O}_2} \) = 4.48 L

d. 50.0 g Ne × \( \frac{1 \text{ mole Ne}}{20.2 \text{ g Ne}} \) × \( \frac{22.4 \text{ L}}{1 \text{ mole Ne}} \) × \( \frac{1000 \text{ mL}}{1 \text{ L}} \) = 55 400 mL

8.54 At STP, the molar volume of any gas is 22.4 L per mole.

a. 2.5 mole N\(_2\) × \( \frac{22.4 \text{ L}}{1 \text{ mole N}_2} \) = 56 L N\(_2\)

b. 0.420 mole He × \( \frac{22.4 \text{ L}}{1 \text{ mole He}} \) × \( \frac{1000 \text{ mL}}{1 \text{ L}} \) = 9410 mL

c. 11.2 L × \( \frac{1 \text{ mole Ne}}{22.4 \text{ L}} \) × \( \frac{20.2 \text{ g Ne}}{1 \text{ mole Ne}} \) = 10.1 g Ne

d. 1600 mL × \( \frac{1 \text{ L}}{1000 \text{ mL}} \) × \( \frac{1 \text{ mole H}_2}{22.4 \text{ L}} \) = 0.071 mole H\(_2\)
8.55  \[ P = \frac{nRT}{V} = \frac{(2.00 \text{ moles})(0.0821 \text{ L \cdot atm})(300 \text{ K})}{(10.0 \text{ L})(\text{mole \cdot K})} = 4.93 \text{ atm} \]

8.56  \[ PV = nRT \]
\[ V = \frac{nRT}{P} = \frac{(4.0 \text{ moles})(0.0821 \text{ L \cdot atm})(291 \text{ K})}{(1.40 \text{ atm})} = 68 \text{ L} \]

8.57  \[ n = \frac{PV}{RT} = \frac{(845 \text{ mm Hg})(20.0 \text{ L})}{(62.4 \text{ L \cdot mm Hg})(295 \text{ K})} \times \frac{32.0 \text{ g O}_2}{1 \text{ mole O}_2} = 29.4 \text{ g O}_2 \]

8.58  \[ 10.0 \text{ g Kr} \times \frac{1 \text{ mole Kr}}{83.8 \text{ g Kr}} = 0.119 \text{ mole Kr} \]
\[ V = \frac{nRT}{P} = \frac{(0.119 \text{ mole})(62.4 \text{ L \cdot mm Hg})(298 \text{ K})}{(575 \text{ mm Hg})} = 3850 \text{ mL} \]

8.59  \[ n = \frac{25.0 \text{ g N}_2}{28.0 \text{ g}} \times \frac{1 \text{ mole N}_2}{1 \text{ mole N}_2} = 0.893 \text{ mole} \]
\[ T = \frac{PV}{nR} = \frac{(630 \text{ mm Hg})(50.0 \text{ L})}{(0.893)(62.4 \text{ L \cdot mm Hg})} = 565 \text{ K} \text{ ( = 292°C)} \]

8.60  \[ n = \frac{0.226 \text{ g CO}_2}{44.0 \text{ g}} \times 1 \text{ mole CO}_2/0.00514 \text{ mole} = 0.000514 \text{ mole} \]
\[ T = \frac{PV}{nR} = \frac{(455 \text{ mm Hg})(0.525 \text{ L})}{(0.00514 \text{ mole})(62.4 \text{ L \cdot mm Hg})} = 745 \text{ K} - 273 = 472 \text{ °C} \]

8.61 Each gas particle in a gas mixture exerts a pressure as it strikes the walls of the container. The total gas pressure for any gaseous sample is thus a sum of all of the individual pressures. When the portion of the pressure due to a particular type of gaseous particle is discussed, it is only part of the total. Accordingly, these “portions” are referred to as “partial” pressures.

8.62 Because the helium and oxygen “particles” in the sample must have the same average kinetic energies, each particle will exert the same average pressure against the container’s walls. For the partial pressures for each of these gases to be same, there must be equal numbers of these two types of gaseous particles.

8.63 To obtain the total pressure in a gaseous mixture, add up all of the partial pressures (provided each carries the same pressure unit).
\[ P_{\text{total}} = P_{\text{Nitrogen}} + P_{\text{Oxygen}} + P_{\text{Helium}} = 425 \text{ torr} + 115 \text{ torr} + 225 \text{ torr} = 765 \text{ torr} \]

8.64 To obtain the total pressure in a gaseous mixture, add up all of the partial pressures, (provided each carries the same pressure unit).
\[ P_{\text{total}} = P_{\text{Argon}} + P_{\text{Neon}} + P_{\text{Nitrogen}} = 415 \text{ mm Hg} + 75 \text{ mm Hg} + 125 \text{ mm Hg} = 615 \text{ mm Hg} \]
\[ 615 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 0.809 \text{ atm} \]
8.65 Because the total pressure in a gaseous mixture is the sum of the partial pressures (provided each carries the same pressure unit), addition and subtraction is used to obtain the “missing” partial pressure.

\[ P_{\text{Nitrogen}} = P_{\text{total}} - (P_{\text{Oxygen}} + P_{\text{Helium}}) \]
\[ = 925 \text{ torr} - (425 \text{ torr} + 75 \text{ torr}) = 425 \text{ torr} \]

8.66 Because the total pressure in a gaseous mixture is the sum of the partial pressures, (provided each carries the same pressure unit), addition and subtraction is used to obtain the “missing” partial pressure.

\[ P_{\text{Helium}} = P_{\text{total}} - P_{\text{Mixture of oxygen, Nitrogen, and Neon}} \]
\[ = 1.50 \text{ atm} - 1.20 \text{ atm} = 0.30 \text{ atm} \]

8.67 a. If oxygen cannot readily cross from the lungs into the bloodstream, then the partial pressure of oxygen will be lower in the blood of an emphysema patient.

b. An increase in the partial pressure of oxygen in the air supplied to the lungs will result in an increase in the partial pressure of oxygen in the bloodstream (addition of reactant causes the formation of more product). Because an emphysema patient has a lower partial pressure of oxygen in the blood, the use of a portable oxygen tank helps to bring the oxygenation of the patient’s blood to a more desirable level.

8.68 a. When ventilation is difficult, the body’s cells are consuming oxygen, but the blood is not being replenished with oxygen in the lungs. Thus, the partial pressure of oxygen in the bloodstream of a person who cannot ventilate properly would be lower than normal. In a similar fashion, the partial pressure of carbon dioxide in the bloodstream would be higher than normal as the cells produce CO\(_2\) but it is not eliminated at the lungs.

b. More gases will pass through the lungs and into the blood when higher gas pressure is introduced by a ventilator. Conversely, more gases will pass out of the blood and into the lungs when a lower gas pressure is introduced into the lungs. Because the air mixture has a higher partial pressure of oxygen than the blood, oxygen would enter the bloodstream. Similarly, the air mixture has a lower partial pressure of carbon dioxide than the blood, and CO\(_2\) would be expelled.

8.69 Gas particles move faster at higher temperatures. This causes the force that each particle exerts against the wall to increase. Pressure is the sum of these forces. Thus, pressure increases with an increase in temperature at constant volume.

8.70 The pressure above the liquid inside the straw is less than the pressure above the liquid elsewhere. The liquid will rise up the straw because it is being pushed harder elsewhere than it is in the straw.

8.71 a. The volume of the chest and lungs will decrease when compressed during the Heimlich maneuver.

b. A decrease in volume causes the pressure to increase. A piece of food would be dislodged with a sufficiently high pressure.

8.72 a. \[ 650 \text{ mm Hg total} \times \frac{21 \text{ mm Hg O}_2}{100. \text{ mm Hg total}} = 140 \text{ mm Hg} \]

b. \[ 100 \text{ mm Hg O}_2 \times \frac{100. \text{ mm Hg total}}{21 \text{ mm Hg O}_2} = 476 \text{ mm Hg} \]

8.73 Recall that all temperatures *must* be in kelvins in computations involving gas laws!

\[ \frac{150 \text{ lb/\text{in.}^2}}{298 \text{ K}} \times \frac{348 \text{ K}}{1 \text{ atm}} \times \frac{1 \text{ atm}}{14.7 \text{ lb/\text{in.}^2}} = 12 \text{ atm} \]


8.74 Remember to convert temperatures to Kelvin units and to use pressure conversion factors for unit cancellation!

\[
750 \text{ L} \times \frac{380 \text{ torr}}{0.20 \text{ atm}} \times \frac{1 \text{ atm}}{760 \text{ torr}} \times \frac{228 \text{ K}}{281 \text{ K}} = 1500 \text{ L}
\]

8.75 Remember to use Kelvin temperature units in the calculation and convert to Celsius degrees after completing the calculation!

\[
400. \text{ K} \times \frac{0.25 \text{ atm}}{2.00 \text{ atm}} = 50. \text{ K} \quad 50. \text{ K} - 273 = -223^\circ \text{C}
\]

8.76 a. \( P_{\text{Nitrogen}} = P_{\text{total}} + VP_{\text{water}} \)

\[
= 745 \text{ mm Hg} - 32 \text{ mm Hg} = 713 \text{ mm Hg}
\]

b. \( T_1 = 30^\circ \text{C} + 273 = 303 \text{ K}; \quad P_1 = 713 \text{ mm Hg}; \quad V_1 = 250 \text{ mL} \)

\( T_2 = 0^\circ \text{C} + 273 = 273 \text{ K}; \quad P_2 = 760 \text{ mm Hg} \)

\[
V_2 = V_1 \times \frac{P_1}{P_2} \times \frac{T_2}{T_1} = 210 \text{ mL}
\]

8.77 The partial pressure of each gas is proportional to the number of particles of each type of gas that is present. Thus, a ratio of partial pressure to total pressure is equal to the ratio of moles of that gas to the total number of moles of gases that are present:

\[
P_{\text{Helium}}/P_{\text{total}} = n_{\text{Helium}}/n_{\text{total}}
\]

Solving the equation for the partial pressure of helium yields:

\[
P_{\text{Helium}} = 2400 \text{ torr} \times \frac{2.0 \text{ mole}}{8.0 \text{ mole}} = 600 \text{ torr} \quad \text{(or } 6.0 \times 10^2 \text{ torr to show two sig figs )}
\]

And for oxygen:

\[
P_{\text{Oxygen}} = 2400 \text{ torr} \times \frac{6.0 \text{ mole}}{8.0 \text{ mole}} = 1800 \text{ torr}
\]

8.78 Because the total pressure is to be reported in mm Hg, the atm and torr units (for argon and nitrogen, respectively) must be converted to mm Hg, as follows:

\[
0.25 \text{ atm} \times \frac{760 \text{ mm Hg}}{1 \text{ atm}} = 190 \text{ mm Hg}
\]

\[
360 \text{ torr} \times \frac{1 \text{ mm Hg}}{1 \text{ torr}} = 360 \text{ mm Hg}
\]

and \( P_{\text{total}} = P_{\text{Argon}} + P_{\text{Helium}} + P_{\text{Nitrogen}} = 190 \text{ mm Hg} + 350 \text{ mm Hg} + 360 \text{ mm Hg} = 900 \text{ mm Hg} \)

8.79 Because the partial pressure of nitrogen is to be reported in torr, the atm and mm Hg units (for oxygen and argon, respectively) must be converted to torr, as follows:

\[
0.60 \text{ atm} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 460 \text{ torr}
\]

\[
425 \text{ mm Hg} \times \frac{1 \text{ torr}}{1 \text{ mm Hg}} = 425 \text{ torr}
\]

and \( P_{\text{Nitrogen}} = P_{\text{total}} - (P_{\text{Oxygen}} + P_{\text{Argon}}) = 1250 \text{ torr} - (460 \text{ torr} + 425 \text{ torr}) = 370 \text{ torr} \)
Because we need the answer in mm Hg we must include a pressure unit conversion factor for oxygen.

\[ P_{\text{oxygen}} = 0.450 \text{ atm} \times \frac{760 \text{ mm Hg}}{1 \text{ atm}} = 342 \text{ mm Hg} \]

255 torr \times \frac{1 \text{ mm Hg}}{1 \text{ torr}} = 255 \text{ mm Hg}

Total pressure = 255 mm Hg + 342 mm Hg = 597 mm Hg

If the initial volume is 1 V, the final volume is \( \frac{1}{0.500} V \) or 0.500 \( V \). (Temperature remains constant).

\[ P_2 = 597 \text{ mm Hg} \times \frac{1 V}{0.500 V} = 1190 \text{ mm Hg} \]

\[ D = \frac{\text{mass}}{\text{volume}} = \frac{32.0 \text{ g O}_2}{1 \text{ mole}} \times \frac{1 \text{ mole}}{22.4 \text{ L (STP)}} = 1.43 \text{ g/L} \]

When a balloon is filled with helium and placed to a hot light, the helium will expand. If the expansion is greater than what the balloon’s material can stretch to accommodate, it will rupture.

a. The partial pressure of oxygen in the lungs is greater than what is present in blood in the alveoli.

b. Arterial blood picks up oxygen in the lungs and delivers oxygen to the body tissues. Arterial blood has a higher partial pressure of oxygen than venous blood.

c. Because carbon dioxide is generated in body tissues, the partial pressure of CO2 will be greater in the tissues than in arterial blood.

d. The venous blood is returning to the lungs from body tissues. Venous blood, which has a higher partial pressure of carbon dioxide than the lungs, transports the excess CO2 to the lungs to be exhaled.

a. The oxygen will diffuse from the lungs into the blood in the alveoli.

b. The oxygen in arterial blood will diffuse into venous blood.

c. The carbon dioxide will diffuse from the tissues into the arterial blood.

d. The carbon dioxide in venous blood will diffuse into the lungs.

The mole-mole conversion factor is obtained from the reaction and used to convert to moles of gas, and the STP molar volume conversion factor is used to convert moles of gas into liters of gas, as shown below:

\[ 2.00 \text{ mole CaCO}_3 \times \frac{1 \text{ mole CO}_2}{1 \text{ mole CaCO}_3} \times \frac{22.4 \text{ L}}{1 \text{ mole CO}_2} = 44.8 \text{ L CO}_2 \]

a. 16.00 g O2 \times \frac{1 \text{ mole O}_2}{32.0 \text{ g}} = 0.500 \text{ mole O}_2

\[ V = \frac{nRT}{P} = \frac{(0.500 \text{ mole})(0.0821 \frac{\text{L atm}}{\text{mole K}})(400 \text{ K})}{(0.800 \text{ atm})(273 \text{ K})} = 20.5 \text{ L} \]

Alternatively, the volume at STP can be calculated for new conditions.

AT STP, the volume of 0.500 mole O2 is

\[ 0.500 \text{ mole O}_2 \times \frac{22.4 \text{ L}}{1 \text{ mole O}_2} = 11.2 \text{ L} \]

Solving for the final volume at 0.800 atm and 400.K

\[ 11.2 \text{ L} \times \frac{400 \text{ K}}{273 \text{ K}} \times \frac{1.00 \text{ atm}}{0.800 \text{ atm}} = 20.5 \text{ L} \]
Chapter 8 Answers and Solutions

b. \[16.00 \, g \, O_2 \times \frac{1 \, \text{mole} \, O_2}{32.0 \, g \, O_2} \times \frac{2 \, \text{mole} \, H_2O}{1 \, \text{mole} \, O_2} = 1.00 \, \text{mole} \, H_2O\]

c. \[V = \frac{nRT}{P} = \frac{(1.00 \, \text{mole})(62.4 \, \text{L} \cdot \text{mm Hg})(388 \, \text{K})}{(760 \, \text{mm Hg})(\text{mole} \cdot \text{K})} = 31.9 \, \text{L}\]

Alternatively, 1.00 mole \(H_2O\) has a volume of 22.4 L at STP, which can be used to calculate the volume at the new conditions when only the temperature changes.

\[22.4 \, \text{L} \times \frac{388 \, \text{K}}{273 \, \text{K}} = 31.8 \, \text{L}\]

8.87 \[8.0 \, g \, \text{Mg} \times \frac{1 \, \text{mole} \, \text{Mg}}{24.3 \, g \, \text{Mg}} \times \frac{1 \, \text{mole} \, O_2}{2 \, \text{mole} \, \text{Mg}} \times 22.4 \, \text{L} \, \text{O}_2 \, (\text{STP}) = 3.7 \, \text{L} \, \text{O}_2\]

8.88 \(n(N_2) = \frac{PV}{RT} = \frac{(840 \, \text{mm Hg})(2.0 \, \text{L})}{(62.4 \, \text{L} \cdot \text{mm Hg})(297 \, \text{K})} \times \frac{2 \, \text{mole} \, \text{NO}_2}{1 \, \text{mole} \, N_2} \times \frac{46.0 \, g \, \text{NO}_2}{1 \, \text{mole} \, \text{NO}_2} = 8.3 \, g \, \text{NO}_2\]

8.89 \(n = \frac{PV}{RT} = \frac{(1.2 \, \text{atm})(35.0 \, \text{L})}{(0.0821 \, \text{L} \cdot \text{atm})(278 \, \text{K})} \times \frac{6.02 \times 10^{23} \, \text{molecules}}{1 \, \text{mole} \, \text{CO}_2} = 1.1 \times 10^{24} \, \text{CO}_2 \, \text{molecules}\)

8.90 \[50.0 \, g \, \text{N}_2 \times \frac{1 \, \text{mole} \, \text{N}_2}{28.0 \, g \, \text{N}_2} = 1.79 \, \text{moles} \, \text{N}_2\]

\[P = \frac{nRT}{V} = \frac{(1.79 \, \text{moles} \, \text{N}_2)(0.0821 \, \text{L} \cdot \text{atm})(298 \, \text{K})}{(15.0 \, \text{L})} = 2.92 \, \text{atm}\]

8.91 \(n = \frac{PV}{RT} = \frac{(2500 \, \text{mm Hg})(2.00 \, \text{L})}{(62.4 \, \text{L} \cdot \text{mm Hg})(291 \, \text{K})} \times \frac{16.0 \, g \, \text{CH}_4}{1 \, \text{mole} \, \text{CH}_4} = 4.4 \, g \, \text{CH}_4\]

8.92 \[4.0 \times 19^{22} \, \text{O}_2 \, \text{molecules} \times \frac{1 \, \text{mole} \, \text{O}_2}{6.02 \times 10^{23} \, \text{molecules}} = 0.066 \, \text{mole} \, \text{O}_2\]

\[V = \frac{nRT}{P} = \frac{(0.066 \, \text{mole})(62.4 \, \text{L} \cdot \text{mm Hg})(278 \, \text{K})}{(845 \, \text{mm Hg})(\text{mole} \cdot \text{K})} = 1.4 \, \text{L} \times \frac{1000 \, \text{mL}}{1 \, \text{L}} = 1400 \, \text{mL}\]

8.93 \[425 \, \text{mL} \times \frac{745 \, \text{mm Hg}}{0.115 \, \text{atm}} \times \frac{1 \, \text{atm}}{760 \, \text{mm Hg}} \times \frac{178 \, \text{K}}{297 \, \text{K}} = 2170 \, \text{mL}\]
8.94 225 mL \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1 \text{ mole gas}}{22.4 \text{ L (STP)}} = 0.0100 \text{ moles of gas}

\text{molar mass} = \frac{\text{g/mole}}{0.0100 \text{ mole gas}} = 115 \text{ g/mole}

\text{Using ideal gas law:} \quad \frac{PV}{MM} = \frac{(g/MM)RT}{PV}

\text{PV} \quad \frac{MM}{PV} = (1.15 \text{ g}) \left( \frac{0.0821 \text{ L \cdot atm}}{\text{mm Hg}} \right) (273 \text{ K}) = 115 \text{ g/mole}

8.95 1 \text{ torr} = 1 \text{ mm Hg} \quad \text{molar mass} = \frac{\text{g of gas}}{\text{moles of gas}}

\text{n} = \frac{PV}{RT} = \frac{(748 \text{ torr})(0.941)}{(62.4 \text{ L \cdot mm Hg})(293 \text{ K})} = 0.0385 \text{ mole}

\text{molar mass} = \frac{1.62 \text{ g}}{0.0385 \text{ mole}} = 42.1 \text{ g/mole}

8.96 762 mL \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1 \text{ mole gas}}{22.4 \text{ L (STP)}} = 0.0340 \text{ moles of gas}

\text{molar mass} = \frac{\text{g/mole}}{0.0340 \text{ mole gas}} = 30. \text{ g/mole}

\text{Using ideal gas law:} \quad \frac{PV}{MM} = \frac{(g/MM)RT}{PV}

\text{MM} = \frac{(1.02 \text{ g})(0.0821 \text{ L \cdot atm}) (273 \text{ K})}{(1.0 \text{ atm})(0.762 \text{ L})(\text{mole} \cdot \text{K})} = 30. \text{ g/mole}

Because \text{CH}_3\text{H}_6\text{ has a molar mass of} 15, \text{there must be two CH}_3\text{ units, which gives a molecular formula of C}_2\text{H}_6\text{.}

8.97 25.0 \text{ g Zn} \times \frac{1 \text{ mole Zn}}{65.4 \text{ g Zn}} \times \frac{1 \text{ mole H}_2}{1 \text{ mole Zn}} \times \frac{22.4 \text{ L H}_2 (\text{STP})}{1 \text{ mole H}_2} = 8.56 \text{ L of H}_2 (\text{g) STP}

8.98 12.0 \text{ g Mg} \times \frac{1 \text{ mole Mg}}{24.3 \text{ g Mg}} \times \frac{1 \text{ mole H}_2}{1 \text{ mole Mg}} = 0.494 \text{ mole H}_2

\text{V} = \frac{nRT}{P} \quad \frac{(0.494 \text{ mole})(0.0821 \text{ L \cdot atm})(297 \text{ K})}{(1.10 \text{ atm})(\text{mole} \cdot \text{K})} = 11.0 \text{ L}

8.99 \text{ a.} 2.5 \times 10^{23} \text{ molecules NO}_2 \times \frac{1 \text{ mole NO}_2}{6.02 \times 10^{23} \text{ molecules}} \times \frac{7 \text{ mole O}_2}{4 \text{ moles NO}_2} \times \frac{22.4 \text{ L O}_2 (\text{STP})}{1 \text{ mole O}_2} \quad = 16 \text{ L O}_2 (\text{g})

\text{ b. Moles (NH}_3\text{)} = \frac{PV}{RT} \quad \frac{(725 \text{ mm Hg})(5.00 \text{ L})}{(62.4 \text{ L \cdot mm Hg})(648 \text{ K})} \times \frac{4 \text{ moles NH}_3}{6 \text{ moles H}_2\text{O}} \times \frac{17.0 \text{ g NH}_3}{1 \text{ mole NH}_3}

\quad = 1.02 \text{ g NH}_3 (\text{g})
8.100  

\[ 1.00 \text{ g CO}_2 \times \frac{1 \text{ mole CO}_2}{44.0 \text{ g CO}_2} = 0.0227 \text{ mole CO}_2 \]

\[ P = \frac{nRT}{V} = \frac{(0.0227 \text{ mole CO}_2)(62.4 \text{ L mm Hg})(297 \text{ K})}{(4.6 \text{ L })(\text{mole} \cdot \text{K})} = 91 \text{ mm Hg} \]

8.101  

\[ 5.4 \text{ g Al} \times \frac{1 \text{ mole Al}}{27.0 \text{ g Al}} \times \frac{3 \text{ mole O}_2}{4 \text{ mole Al}} \times \frac{22.4 \text{ L O}_2 \text{ (STP)}}{1 \text{ mole H}_2} = 3.4 \text{ L O}_2 \text{ (g)} \]