CHAPTER 5 GASES AND THE KINETIC-MOLECULAR THEORY

5.1   a) The volume of the liquid remains constant, but the volume of the gas increases to the volume of the larger container.
b) The volume of the container holding the gas sample increases when heated, but the volume of the container holding the liquid sample remains essentially constant when heated.
c) The volume of the liquid remains essentially constant, but the volume of the gas is reduced.

5.2   The particles in a gas are further apart than those are in a liquid.
a) The greater empty space between gas molecules allows gases to be more compressible than liquids.
b) The greater empty space between gas molecules allows gases to flow with less resistance (hindrance) than liquids.
c) The large empty space between gas molecules limits their interaction, allowing all mixtures of gases to be solutions.
d) The large empty space between gas molecules increases the volume of the gas, therefore decreasing the density.

5.3   The mercury column in the mercury barometer stays up due to the force exerted by the atmosphere on the mercury in the outer reservoir just balancing the gravitational force on the mercury in the tube. Its height adjusts according to the air pressure on the reservoir. On a mountaintop, the air pressure is less, so the height of mercury it balances is shorter than at sea level.

5.4   The pressure of mercury is its weight (force) per unit area. The weight, and thus the pressure, of the mercury column is directly proportional to its height.

5.5   When the mercury level in the arm attached to the flask is higher than the level in the other arm, the pressure in the flask is less than the pressure exerted in the other arm. This is an impossible situation for a closed-end manometer as the flask pressure cannot be less than the vacuum in the other arm.

5.6   The ratio of the heights of columns of mercury and water are inversely proportional to the ratio of the densities of the two liquids.

$$\frac{h_{H_2O}}{h_{Hg}} = \frac{d_{Hg}}{d_{H_2O}}$$

$$h_{H_2O} = \frac{d_{Hg}}{d_{H_2O}} x h_{Hg} = \left(\frac{13.5 \text{ g/mL}}{1.00 \text{ g/mL}}\right)(725 \text{ mmHg})\left(\frac{10^{-3} \text{ m}}{1 \text{ mm}}\right)\left(\frac{1 \text{ cm}}{10^{-2} \text{ m}}\right) = 978.75 = \textbf{979 cm H}_2\textbf{O}$$

5.7 $P (\text{mmH}_2\text{O}) = h_{H_2O} = \frac{d_{Hg}}{d_{H_2O}} x h_{Hg} = \left(\frac{13.5 \text{ g/mL}}{1.00 \text{ g/mL}}\right)(755 \text{ mmHg})\left(\frac{10^{-3} \text{ m}}{1 \text{ mm}}\right)\left(\frac{1 \text{ cm}}{10^{-2} \text{ m}}\right) = 1019.25 = \textbf{1.02 x 10}^3 \textbf{ cm H}_2\textbf{O}$

5.8   a) $(0.745 \text{ atm})\left(\frac{760 \text{ mmHg}}{1 \text{ atm}}\right) = 566.2 = \textbf{566 mmHg}$

b) $(992 \text{ torr})\left(\frac{1.01325 \text{ bar}}{760 \text{ torr}}\right) = 1.32256 = \textbf{1.32 bar}$

c) $(365 \text{ kPa})\left(\frac{1 \text{ atm}}{101.325 \text{ kPa}}\right) = 3.60227 = \textbf{3.60 atm}$

d) $(804 \text{ mmHg})\left(\dfrac{101.325 \text{ kPa}}{760 \text{ mmHg}}\right) = 107.191 = \textbf{107 kPa}$

5.9 a) $(74.8 \text{ cmHg})\left(\dfrac{10^{-2} \text{ m}}{1 \text{ cm}}\right)\left(\dfrac{1 \text{ mm}}{10^{-3} \text{ m}}\right)\left(\dfrac{1 \text{ atm}}{760 \text{ mmHg}}\right) = 0.98421 = \textbf{0.984 atm}$

b) $(27.0 \text{ atm})\left(\dfrac{101.325 \text{ kPa}}{1 \text{ atm}}\right) = 2735.775 = \textbf{2.74 x } 10^3 \textbf{ kPa}$

c) $(8.50 \text{ atm})\left(\dfrac{1.01325 \text{ bar}}{1 \text{ atm}}\right) = 8.6126 = \textbf{8.61 bar}$

d) $(0.907 \text{ kPa})\left(\dfrac{760 \text{ torr}}{101.325 \text{ kPa}}\right) = 6.80306 = \textbf{6.80 torr}$

5.10 Since the height of the mercury column in contact with the gas is higher than the column in contact with the air, the gas is exerting less pressure on the mercury than the air.

$(2.35 \text{ cm})\left(\dfrac{10^{-2} \text{ m}}{1 \text{ cm}}\right)\left(\dfrac{1 \text{ mm}}{10^{-3} \text{ m}}\right)\left(\dfrac{1 \text{ torr}}{1 \text{ mmHg}}\right) = 23.5 \text{ torr}$

738.5 torr - 23.5 torr = 715.0 torr

$\text{P(atm)} = (715.0 \text{ torr})\left(\dfrac{1 \text{ atm}}{760 \text{ torr}}\right) = 0.940789 = \textbf{0.9408 atm}$

5.11 Since the height of the mercury column in contact with the gas is higher than the column in contact with the air, the gas is exerting less pressure on the mercury than the air.

$(1.30 \text{ cm})\left(\dfrac{10^{-2} \text{ m}}{1 \text{ cm}}\right)\left(\dfrac{1 \text{ mm}}{10^{-3} \text{ m}}\right)\left(\dfrac{1 \text{ torr}}{1 \text{ mmHg}}\right) = 13.0 \text{ torr}$

765.2 torr - 13.0 torr = 752.2 torr

$\text{P(kPa)} = (752.2 \text{ torr})\left(\dfrac{101.325 \text{ kPa}}{760 \text{ torr}}\right) = 100.285 = \textbf{100.3 kPa}$

5.12 The difference in the height of the Hg is directly related to pressure in atmospheres.

$\text{P(atm)} = (0.734 \text{ mHg})\left(\dfrac{1 \text{ mmHg}}{10^{-3} \text{ mHg}}\right)\left(\dfrac{1 \text{ atm}}{760 \text{ mmHg}}\right) = 0.965789 = \textbf{0.966 atm}$

5.13 $\text{P(Pa)} = (3.56 \text{ cm})\left(\dfrac{10^{-2} \text{ mHg}}{1 \text{ cmHg}}\right)\left(\dfrac{1 \text{ mmHg}}{10^{-3} \text{ m}}\right)\left(\dfrac{1.01325 \text{ x } 10^5 \text{ Pa}}{760 \text{ mmHg}}\right) = 4746.276 = \textbf{4.75 x } 10^3 \textbf{ Pa}$

5.14 a) $\text{P(atm)} = (2.75 \text{ x } 10^2 \text{ mmHg})\left(\dfrac{1 \text{ atm}}{760 \text{ mmHg}}\right) = 0.361842 = \textbf{0.362 atm}$

b) $\text{P(atm)} = (91 \text{ psi})\left(\dfrac{1 \text{ atm}}{14.7 \text{ psi}}\right) = 6.190 = \textbf{6.2 atm}$

c) $\text{P(atm)} = (9.15 \text{ x } 10^6 \text{ Pa})\left(\dfrac{1 \text{ atm}}{1.01325 \text{ x } 10^5 \text{ Pa}}\right) = 90.303 = \textbf{90.3 atm}$

d) $P(atm) = \left(2.44 \times 10^4 \text{ torr}\right)\left(\dfrac{1 \text{ atm}}{760 \text{ torr}}\right) = 32.105 = \textbf{32.1 atm}$

5.15   a) 1 atm = $1.01325 \times 10^5 \text{ N/m}^2$
The force on 1 m^2 of ocean is $1.01325 \times 10^5 \text{ N}$.
$F = m \times g$
$1.01325 \times 10^5 \text{ N} = mg$
$1.01325 \times 10^5 \dfrac{kg \cdot m}{s^2} = (mass)(9.81 \text{ m/s}^2)$
      mass $= 1.03287 \times 10^4 = \textbf{1.03} \times \textbf{10}^\textbf{4} \textbf{ kg}$

b) $\left(1.03287 \times 10^4 \dfrac{kg}{m^2}\right)\left(\dfrac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\dfrac{10^{-2} \text{ m}}{1 \text{ cm}}\right)^2 = 1.03287 \times 10^3 \text{ g/cm}^2$ (unrounded)

       Height $= \left(1.03287 \times 10^3 \dfrac{g}{cm^2}\right)\left(\dfrac{1 \text{ mL}}{22.6 \text{ g}}\right)\left(\dfrac{1 \text{ cm}^3}{1 \text{ mL}}\right) = 45.702 = \textbf{45.7 cm Os}$

5.16   The statement is incomplete with respect to temperature and mass of sample. At constant temperature and moles of gas, the volume of gas is inversely proportional to the pressure.

5.17   variable   fixed
         a) Volume; Temperature   Pressure; Moles
         b) Moles; Volume   Temperature; Pressure
         c) Pressure; Temperature   Volume; Moles

5.18   At constant temperature and volume, the pressure of the gas is directly proportional to the number of moles of the gas. Verify this by examining the ideal gas equation. At constant T and V, the ideal gas equation becomes P = n(RT/V) or P = n x constant.

5.19   a) **n fixed**   b) **P halved**   c) **P fixed**   d) **T doubled**

5.20   a) As the pressure on a gas increases, the molecules move closer together, decreasing the volume. When the pressure is tripled, the **volume decreases to one third of the original volume** at constant temperature (Boyle's Law).
b) As the temperature of a gas increases, the gas molecules gain kinetic energy. With higher energy, the gas molecules collide with the walls of the container with greater force, which increases the size (volume) of the container. If the temperature is increased by a factor of 2.5 (at constant pressure) then the **volume will increase by a factor of 2.5** (Charles's Law).
c) As the number of molecules of gas increase, the force they exert on the container increases. This results in an increase in the volume of the container. Adding two moles of gas to one mole increases the number of moles by a factor of three, thus the **volume increases by a factor of three** (Avogadro's Law).

5.21 $V_2 = V_1 \times (P_1/P_2) \times (T_2/T_1)$
a) If P_2 is 1/4 of P_1, then the volume would be **increased by a factor of 4**.
b) $V_2 = V_1 \times (101 \text{ kPa}/202 \text{ kPa}) \times (155 \text{ K}/310 \text{ K}) = 1/4$, so the volume would be **decreased by a factor of 4**.
c) $V_2 = V_1 \times (202 \text{ kPa}/101 \text{ kPa}) \times (305 \text{ K}/305 \text{ K}) = 2$, so the volume would be **increased by a factor of 2**.

5.22   a) The temperature is decreased by a factor of 2, so **the volume is decreased by a factor of 2** (Charles's Law).
b) The temperature increases by a factor of $[(700 + 273) / (350 + 273)] = 1.56$, so **the volume is increased by a factor of 1.56** (Charles's Law).
c) The pressure is increased by a factor of 4, so **the volume decreases by a factor of 4** (Boyle's Law).

5.23 $V_2 = V_1 \times (P_1/P_2) \times (T_2/T_1) \times (n_2/n_1)$
a) Since $n_2 = 0.5\ n_1$, the **volume would be decreased by a factor of 2**.
b) This corresponds to both P and T remaining constant, so **the volume remains constant**.
c) Since $P_2 = 0.25\ P_1$ and $T_2 = 0.25\ T_1$, these two effects offset one another and **the volume remains constant**.

5.24   <u>Plan:</u> The temperature must be lowered to reduce the volume of a gas. Charles's Law states that at constant pressure and with a fixed amount of gas, the volume of a gas is directly proportional to the absolute temperature of the gas.
<u>Solution:</u>
$(V_1 / T_1) = (V_2 / T_2)$ at constant n and P
$T_2 = T_1 (V_2 / V_1)$

$$T_2 = \left((273 + 198)\,K\right)\left(\frac{2.50\ L}{5.10\ L}\right) - 273 = -42.1167 = \textbf{-42°C}$$

5.25 $V_2 = V_1 (T_2 / T_1) = \left(93\ L\right)\left(\dfrac{(273 - 22)\,K}{(273 + 145)\,K}\right) = 55.844 = \textbf{56 L}$

5.26   <u>Plan:</u> To find V_2, summarize and convert gas variables and apply the combined gas law equation.
$(P_1V_1/T_1 = P_2V_2T_2)$
<u>Solution:</u>

$$V_2 = V_1 (T_2 / T_1) (P_1 / P_2) = \left(25.5\ L\right)\left(\frac{273\ K}{298\ K}\right)\left(\frac{153.3\ kPa}{101.325\ kPa}\right) = 35.3437 = \textbf{35.3 L}$$

5.27 $V_2 = V_1 (T_2 / T_1) (P_1 / P_2) = \left(3.65\ L\right)\left(\dfrac{(273 - 14)\,K}{298\ K}\right)\left(\dfrac{745\ torr}{367\ torr}\right) = 6.4397 = \textbf{6.44 L}$

5.28   <u>Plan:</u> Given the volume, pressure, and temperature of a gas, the number of moles of the gas can be calculated using the ideal gas equation, $n = PV/RT$. The gas constant, $R = 0.0821$ L•atm/mol•K, gives pressure in atmospheres and temperature in Kelvin. The given pressure in torr must be converted to atmospheres and the temperature converted to Kelvin.
<u>Solution:</u>

$$n = PV / RT = \frac{(228\ torr)(5.0\ L)}{\left(0.0821\dfrac{L \cdot atm}{mol \cdot K}\right)\left((273 + 27)\,K\right)}\left(\frac{1\ atm}{760\ torr}\right) = 0.060901 = \textbf{0.061 mol chlorine}$$

5.29 $PV = nRT$

$$P = nRT / V = \frac{\left(1.47 \times 10^{-3}\ mol\right)\left(0.0821\dfrac{L \cdot atm}{mol \cdot K}\right)\left((273 + 26)\,K\right)}{75.0\ mL}\left(\frac{1\ mL}{10^{-3}\ L}\right)\left(\frac{760\ torr}{1\ atm}\right)$$

        $= 365.6655 = \textbf{366 torr}$

5.30   Plan: Solve the ideal gas equation for moles and convert to mass using the molar mass of ClF_3.
        Solution:
        $PV = (m / \mathcal{M})RT$

$$m = PV\,\mathcal{M}/RT = \frac{(699 \text{ mmHg})(207 \text{ mL})\left(92.45\dfrac{g}{mol}\right)}{\left(0.0821\dfrac{L \cdot atm}{mol \cdot K}\right)((273 + 45)K)}\left(\frac{1 \text{ atm}}{760 \text{ mmHg}}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)$$

$$= 0.67417 = \textbf{0.674 g ClF}_3$$

5.31 $PV = (m/\mathcal{M})RT$

$$P = mRT/\mathcal{M}\,V = \frac{(75.0 \text{ g})\left(0.0821\dfrac{L \cdot atm}{mol \cdot K}\right)((273 + 115)K)}{\left(44.02\dfrac{g}{mol}\right)(3.1 \text{ L})} = 17.5075 = \textbf{18 atm N}_2\textbf{O}$$

5.32 $PV = nRT$ The total pressure of the gas is $(85 + 14.7)$psi.

$$n = PV / RT = \frac{((85 + 14.7)\text{psi})(1.5 \text{ L})}{\left(0.0821\dfrac{L \cdot atm}{mol \cdot K}\right)((273 + 23)K)}\left(\frac{1 \text{ atm}}{14.7 \text{ psi}}\right) = 0.41863 = \textbf{0.42 mol SO}_2$$

5.33   Plan: Assuming that while rising in the atmosphere the balloon will neither gain nor lose gas molecules, the
        number of moles of gas calculated at sea level will be the same as the number of moles of gas at the higher
        altitude. Using the ideal gas equation, (n x R) is a constant equal to (PV/T). Given the sea-level conditions of
        volume, pressure and temperature, and the temperature and pressure at the higher altitude for the gas in the
        balloon, we can set up an equation to solve for the volume at the higher altitude. Comparing the calculated volume
        to the given maximum volume of 835 L will tell us if the balloon has reached its maximum volume at this altitude.
        Solution:
        $(P_1V_1 / T_1) = (P_2V_2 / T_2)$

$$V_2 = (P_1V_1T_2 / T_1P_2) = \frac{(755 \text{ torr})(55.0 \text{ L})((273 - 5)K)}{((273 + 23)K)(0.066 \text{ atm})}\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right) = 749.54 = 7.5 \text{ x } 10^2 \text{ L}$$

        The calculated volume of the gas at the higher altitude is less than the maximum volume of the balloon. **No**, the
        balloon will not reach its maximum volume.
        Check: Should we expect that the volume of the gas in the balloon should increase? At the higher altitude, the
        pressure decreases, this increases the volume of the gas. At the higher altitude, the temperature decreases, this
        decreases the volume of the gas. Which of these will dominate? The pressure decreases by a factor of 0.99/0.066
        = 15. If we label the initial volume V_1, then the resulting volume is 15 V_1. The temperature decreases by a factor
        of 296/268 = 1.1, so the resulting volume is V_1/1.1 or 0.91 V_1. The increase in volume due to the change in
        pressure is greater than the decrease in volume due to change in temperature, so the volume of gas at the higher
        altitude should be greater than the volume at sea level.

5.34   Air is mostly N_2 (28.02 g/mol), O_2 (32.00 g/mol), and argon (39.95 g/mol). These "heavy" gases dominate the
        density of dry air. Moist air contains H_2O (18.02 g/mol). The relatively light water molecules lower the density of
        the moist air.

5.35   The molar mass of H_2 is less than the average molar mass of air (mostly N_2, O_2, and Ar), so air is denser. To
        collect a beaker of $H_2(g)$, **invert** the beaker so that the air will be replaced by the lighter H_2. The molar mass of
        CO_2 is greater than the average molar mass of air, so $CO_2(g)$ is more dense. Collect the CO_2 holding the beaker
        upright, so the lighter air will be displaced out the top of the beaker.

5.36   Gases mix to form a solution and each gas in the solution behaves as if it were the only gas present.

5.37 $P_A = X_A P_T$ The partial pressure of a gas (P_A) in a mixture is directly proportional to its mole fraction (X_A).

5.38   <u>Plan:</u> Using the ideal gas equation and the molar mass of xenon, 131.3 g/mol, we can find the density of xenon gas at STP. Standard temperature is 0°C and standard pressure is 1 atm. Do not forget that the pressure at STP is exact and will not affect the significant figures.
<u>Solution:</u>

$$d = \mathcal{M} \, P \, / \, RT = \frac{(131.3 \text{ g/mol})(1 \text{ atm})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(273 \text{ K})} = 5.8581 = \textbf{5.86 g/L}$$

5.39 $d = \mathcal{M} \, P \, / \, RT = \dfrac{(137.86 \text{ g/mol})(1.5 \text{ atm})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\big((273 + 120)\text{K}\big)} = 6.40905 = \textbf{6.4 g/L}$

5.40   <u>Plan:</u> Apply the ideal gas equation to determine the number of moles. Convert moles to mass and divide by the volume to obtain density in g/L. Do not forget that the pressure at STP is exact and will not affect the significant figures.
<u>Solution:</u>

$$n = PV \, / \, RT = \frac{(1 \text{ atm})(0.0400 \text{ L})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(273 \text{ K})} = 1.78465 \times 10^{-3} = \textbf{1.78} \times \textbf{10}^{-3} \textbf{ mol AsH}_3$$

$$d = \text{mass} \, / \, \text{volume} = \frac{(1.78465 \times 10^{-3} \text{ mol})(77.94 \text{ g/mol})}{(0.0400 \text{ L})} = 3.47740 = \textbf{3.48 g/L}$$

5.41 $d = \mathcal{M} \, P \, / \, RT$

$$\mathcal{M} = dRT \, / \, P = \frac{(2.71 \text{ g/L})\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\big((273 + 0)\text{K}\big)}{(3.00 \text{ atm})} = 20.24668 = 20.2 \text{ g/mol}$$

Therefore, the gas is **Ne**.

5.42   <u>Plan:</u> Rearrange the formula PV = (m / $\mathcal{M}$)RT to solve for molar mass: $\mathcal{M}$ = mRT / PV.
<u>Solution:</u>

$$\mathcal{M} = mRT \, / \, PV = \frac{(206 \text{ ng})\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\big((273 + 45)\text{K}\big)}{(388 \text{ torr})(0.206 \text{ μL})}\left(\frac{10^{-9} \text{ g}}{1 \text{ ng}}\right)\left(\frac{1 \text{μL}}{10^{-6}\text{L}}\right)\left(\frac{760 \text{ torr}}{1 \text{ atm}}\right)$$

$$= 51.13899 = \textbf{51.1 g/mol}$$

5.43 $\mathcal{M} = mRT \, / \, PV = \dfrac{(0.103 \text{ g})\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\big((273 + 22)\text{K}\big)}{(747 \text{ mmHg})(63.8 \text{ mL})}\left(\dfrac{1 \text{ mL}}{10^{-3} \text{ L}}\right)\left(\dfrac{760 \text{ mmHg}}{1 \text{ atm}}\right)$

$$= 39.7809 = 39.8 \text{ g/mol}$$

The molar masses are N_2 = 28 g/mol, Ne = 20 g/mol, and Ar = 40 g/mol
Therefore, the gas is **Ar**.

5.44   Plan: Use the ideal gas equation to determine the number of moles of Ar and O_2. The gases are combined ($n_{TOT} = n_{Ar} + n_O$) into a 400 mL flask (V) at 27°C (T). Determine the total pressure from n_{TOT}, V, and T.
Solution:
$n = PV / RT$

$$\text{Moles Ar} = PV / RT = \frac{(1.20 \text{ atm})(0.600 \text{ L})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 227)\text{K})} = 0.017539585 \text{ mol Ar (unrounded)}$$

$$\text{Moles } O_2 = PV / RT = \frac{(501 \text{ torr})(0.200 \text{ L})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 127)\text{K})}\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right) = 0.004014680 \text{ mol } O_2 \text{ (unrounded)}$$

$$\text{P mixture} = nRT / V = \frac{((0.017539585 + 0.004014680)\text{mol})\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 27)\text{K})}{400 \text{ mL}}\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$$

$= 1.32720 = \textbf{1.33 atm}$

5.45 $$\text{Total moles} = n_T = PV/RT = \frac{(626 \text{ mmHg})(355 \text{ mL})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 35)\text{K})}\left(\frac{1 \text{ atm}}{760 \text{ mmHg}}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)$$

$= 0.011563657 \text{ mol (unrounded)}$
Moles Ne = n_{Ne} = (0.146 g Ne) (1 mol Ne / 20.18 g Ne) = 0.007234886 mol Ne (unrounded)
Moles Ar = n_T - n_{Ne} = (0.011563657 - 0.007234886) mol = 0.004328771 = **0.0043 mol Ar**

5.46 $d = \mathcal{M} P / RT$
At 17°C

$$d = \mathcal{M} P / RT = \frac{(28.8 \text{ g/mol})(744 \text{ torr})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 17)\text{K})}\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right) = 1.18416 = \textbf{1.18 g/L}$$

At 60.0°C

$$d = \mathcal{M} P / RT = \frac{(28.8 \text{ g/mol})(744 \text{ torr})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 60.0)\text{K})}\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right) = 1.03125 = \textbf{1.03 g/L}$$

5.47 $PV = nRT$
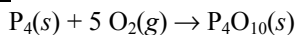
$$n / V = P / RT = \frac{(650. \text{ torr})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 - 25)\text{K})}\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right) = 0.042005 = \textbf{0.0420 mol/L}$$

5.48   Plan: The problem gives the mass, volume, temperature and pressure of a gas, so we can solve for molar mass using $\mathcal{M}$ = mRT/PV. The problem also states that the gas is a hydrocarbon, which by, definition, contains only carbon and hydrogen atoms. We are also told that each molecule of the gas contains five carbon atoms so we can use this information and the calculated molar mass to find out how many hydrogen atoms are present and the formula of the compound.

Solution:

$$\mathscr{M} = mRT/PV = \frac{(0.482 \text{ g})\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\big((273 + 101)\text{K}\big)}{(767 \text{ torr})(0.204 \text{ L})}\left(\frac{760 \text{ torr}}{1 \text{ atm}}\right) = 71.8869 \text{ g/mol (unrounded)}$$

The carbon accounts for [5 (12 g/mol)] = 60 g/mol, thus, the hydrogen must make up the difference (72 - 60) = 12 g/mol. A value of 12 g/mol corresponds to 12 H atoms. (Since fractional atoms are not possible, rounding is acceptable.)
Therefore, the molecular formula is **C_5H_{12}**.

5.49 $PV = nRT$
$n = PV / RT$
Molecules of air = n x Avogadro's number = (PV / RT) x Avogadro's number

$$\text{Molecules of air} = \frac{(1.00 \text{ atm})(1.00 \text{ L})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\big((273 + 25)\text{K}\big)}\left(\frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}}\right)$$

$= 2.461395 \times 10^{22}$ molecules (unrounded)
Molecules N_2 = (2.461395 x 10^{22} molecules) (78.08%/100%) = 1.921857 x 10^{22} = **1.92 x 10^{22} molecules N_2**
Molecules O_2 = (2.461395 x 10^{22} molecules) (20.94%/100%) = 5.154161 x 10^{21} = **5.15 x 10^{21} molecules O_2**
Molecules CO_2 = (2.461395 x 10^{22} molecules) (0.05%/100%) = 1.2306975 x 10^{19} = **1 x 10^{19} molecules O_2**
Molecules Ar = (2.461395 x 10^{22} molecules) (0.93%/100%) = 2.289097 x 10^{20} = **2.3 x 10^{20} molecules O_2**

5.50   a) $PV = nRT$

$$n = PV / RT = \frac{(850. \text{ torr})(21 \text{ L})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\big((273 + 45)\text{K}\big)}\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right) = 0.89961 = \textbf{0.90 mol gas}$$

b) The information given in ppm is a way of expressing the proportion, or fraction, of SO_2 present in the mixture. Since n is directly proportional to V, the *volume* fraction can be used in place of the *mole* fraction used in equation 5.12. There are 7.95 x 10^3 parts SO_2 in a million parts of mixture, so volume fraction = (7.95 x 10^3 / 1 x 10^6) = 7.95 x 10^{-3}.
Therefore, P_{SO_2} = volume fraction x P_{TOT} = (7.95 x 10^{-3}) (850. torr) = 6.7575 = **6.76 torr**.

5.51   Plan: We can find the moles of oxygen from the standard molar volume of gases and use the stoichiometric ratio from the balanced equation to determine the moles of phosphorus that will react with the oxygen.
Solution:
$$P_4(s) + 5 \text{ O}_2(g) \rightarrow P_4O_{10}(s)$$
$$\text{Mass } P_4 = (35.5 \text{ L O}_2)\left(\frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2}\right)\left(\frac{1 \text{ mol P}_4}{5 \text{ mol O}_2}\right)\left(\frac{123.88 \text{ g P}_4}{1 \text{ mol P}_4}\right) = 39.2655 = \textbf{39.3 g P}_4$$
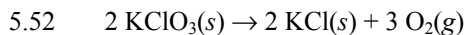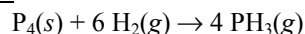
5.52 $2 \text{ KClO}_3(s) \rightarrow 2 \text{ KCl}(s) + 3 \text{ O}_2(g)$
Moles O_2 = $n = PV / RT$
Moles of O_2 x mole ratio x molar mass $KClO_3$ = g $KClO_3$

$$\text{Mass KClO}_3 = \left[\frac{(752 \text{ torr})(638 \text{ mL})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\big((273 + 128)\text{K}\big)}\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\right]\left(\frac{2 \text{ mol KClO}_3}{3 \text{ mol O}_2}\right)\left(\frac{122.55 \text{ g KClO}_3}{1 \text{ mol KClO}_3}\right)$$

$= 1.56660 = \textbf{1.57 g KClO}_3$

5.53   <u>Plan:</u> To find the mass of PH_3, write the balanced equation, convert mass of P_4 to moles, solve for moles of H_2 using the standard molar volume (or use ideal gas equation), and proceed with the limiting reactant problem.
<u>Solution:</u>
$$P_4(s) + 6\ H_2(g) \rightarrow 4\ PH_3(g)$$
Moles hydrogen = (83.0 L) (1 mol / 22.4 L) = 3.705357 mol H_2 (unrounded)
Moles phosphorus = (37.5 g P_4) (1 mol P_4 / 123.88 g P_4) = 0.302712 mol P_4 (unrounded)
Dividing each of the moles by the coefficients shows that P_4 is the limiting reactant.

Mass PH_3 = $\left(37.5\ \text{g}\ P_4\right)\left(\dfrac{1\ \text{mol}\ P_4}{123.88\ \text{g}\ P_4}\right)\left(\dfrac{4\ \text{mol}\ PH_3}{1\ \text{mol}\ P_4}\right)\left(\dfrac{33.99\ \text{g}\ PH_3}{1\ \text{mol}\ PH_3}\right)$ = 41.15676 = **41.2 g PH_3**

5.54 $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(l)$
The moles are directly proportional to the volumes of the gases at the same temperature and pressure. Thus, the limiting reactant may be found by comparing the volumes of the gases. The volumes are divided by the coefficient in the balanced chemical equation, and the lower amount is limiting.
(35.6 L NH_3 / 4) = 8.90 L NH_3
(40.5 L O_2 / 5) = 8.10 L O_2 O_2 is the limiting reactant.

Mass NO = $\left(40.5\ \text{L}\ O_2\right)\left(\dfrac{1\ \text{mol}\ O_2}{22.4\ \text{L}\ O_2}\right)\left(\dfrac{4\ \text{mol}\ NO}{5\ \text{mol}\ O_2}\right)\left(\dfrac{30.01\ \text{g}\ NO}{1\ \text{mol}\ NO}\right)$ = 43.4073 = **43.4 g NO**

5.55   <u>Plan:</u> First, write the balanced equation. The moles of hydrogen produced can be calculated from the ideal gas equation and then the stoichiometric ratio from the balanced equation is used to determine the moles of aluminum that reacted. The problem specifies "hydrogen gas collected over water," so the partial pressure of water must first be subtracted. Table 5.3 reports pressure at 26°C (25.2 torr) and 28°C (28.3 torr), so take the average of the two values.
<u>Solution:</u>
$$2\ Al(s) + 6\ HCl(aq) \rightarrow 2\ AlCl_3(aq) + 3\ H_2(g)$$
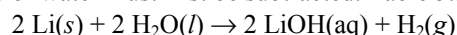Hydrogen pressure = total pressure - pressure from water =
(751 mmHg) - [(28.3 + 25.2) torr / 2] = 724.25 torr (unrounded)
Moles of hydrogen = n = PV / RT
Mass of aluminum = mol of hydrogen x mole ratio x molar mass of Al = PV/RT x mole ratio x molar mass of Al
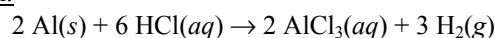
Mass of Al = $\left[\dfrac{\left(724.25\ \text{torr}\right)\left(35.8\ \text{mL}\right)}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left(\left(273 + 27\right)\text{K}\right)}\left(\dfrac{1\ \text{atm}}{760\ \text{torr}}\right)\left(\dfrac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)\right]\left(\dfrac{2\ \text{mol}\ Al}{3\ \text{mol}\ H_2}\right)\left(\dfrac{26.98\ \text{g}\ Al}{1\ \text{mol}\ Al}\right)$

= 0.024914 = **0.0249 g Al**

5.56   First, write the balanced equation. The problem specifies "hydrogen gas collected over water," so the partial pressure of water must first be subtracted. Table 5.3 reports pressure at 18°C (15.52 torr).
$$2\ Li(s) + 2\ H_2O(l) \rightarrow 2\ LiOH(aq) + H_2(g)$$
Pressure H_2 = total pressure - vapor pressure of H_2O = 725 mmHg - 15.5 torr (1mmHg/1torr)
Moles H_2 = g Li x 1/molar mass Li x mole ratio
Volume H_2 = V = nRT / P = (g Li x 1/molar mass Li x mole ratio)RT / P

Volume H_2 = $\dfrac{\left[\left(0.84\ \text{g}\ Li\right)\left(\dfrac{1\ \text{mol}\ Li}{6.941\ \text{g}\ Li}\right)\left(\dfrac{1\ \text{mol}\ H_2}{2\ \text{mol}\ Li}\right)\right]\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left(\left(273 + 18\right)\text{K}\right)}{\left(725 - 15.5\right)\text{mmHg}}\left(\dfrac{760\ \text{mmHg}}{1\ \text{atm}}\right)$

= 1.5485 = **1.5 L H_2**

5.57   <u>Plan:</u> To find mL of SO_2, write the balanced equation, convert the given mass of P_4S_3 to moles, use the molar ratio to find moles of SO_2, and use the ideal gas equation to find volume.
    <u>Solution:</u>

$$P_4S_3(s) + 8\ O_2(g) \rightarrow P_4O_{10}(s) + 3\ SO_2(g)$$
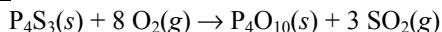
Moles SO_2 = mass P_4S_3 x 1/molar mass P_4S_3 x mole ratio

Volume SO_2 = nRT / P = (mass P_4S_3 x 1/molar mass P_4S_3 x mole ratio)RT / P

$$V = \frac{\left[(0.800\text{ g }P_4S_3)\left(\dfrac{1\text{ mol }P_4S_3}{220.09\text{ g }P_4S_3}\right)\left(\dfrac{3\text{ mol }SO_2}{1\text{ mol }P_4S_3}\right)\right]\left(0.0821\dfrac{L\cdot atm}{mol\cdot K}\right)((273+32)K)}{725\text{ torr}}\left(\frac{760\text{ torr}}{1\text{ atm}}\right)\left(\frac{1\text{ mL}}{10^{-3}\text{ L}}\right)$$

    = 286.239 = **286 mL**

5.58 $CCl_4(g) + 2\ HF(g) \rightarrow CF_2Cl_2(g) + 2\ HCl(g)$

Moles Freon = n = PV / RT

Mass CCl_4 = mol Freon x mole ratio x molar mass of CCl_4 = (PV / RT) x mole ratio x molar mass of CCl_4

$$= \left[\frac{(1.20\text{ atm})(16.0\text{ dm}^3)}{\left(0.0821\dfrac{L\cdot atm}{mol\cdot K}\right)((273+27)K)}\left(\frac{1\text{ L}}{1\text{ dm}^3}\right)\right]\left(\frac{1\text{ mol }CCl_4}{1\text{ mol }CF_2Cl_2}\right)\left(\frac{153.81\text{ g }CCl_4}{1\text{ mol }CCl_4}\right)$$

    = 119.9006 = **1.20 x 10^2 g CCl_4**

5.59   <u>Plan:</u> First, write the balanced equation. Given the amount of xenon hexafluoride that reacts, we can find the number of moles of silicon tetrafluoride gas formed. Then, using the ideal gas equation with the moles of gas, the temperature and the volume, we can calculate the pressure of the silicon tetrafluoride gas.
    <u>Solution:</u>

$$2\ XeF_6(s) + SiO_2(s) \rightarrow 2\ XeOF_4(l) + SiF_4(g)$$
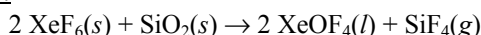
Mole SiF_4 = n = mass XeF_6 x 1/molar mass XeF_6 x mole ratio

Pressure SiF_4 = P = nRT / V = (mass XeF_6 x 1/molar mass XeF_6 x mole ratio)RT / V

$$P = \frac{\left[(2.00\text{ g }XeF_6)\left(\dfrac{1\text{ mol }XeF_6}{245.3\text{ g }XeF_6}\right)\left(\dfrac{1\text{ mol }SiF_4}{2\text{ mol }XeF_6}\right)\right]\left(0.0821\dfrac{L\cdot atm}{mol\cdot K}\right)((273+25)K)}{1.00\text{ L}}$$

    = 0.099738 = **0.0997 atm SiF_4**

5.60 $2\ PbS(s) + 3\ O_2(g) \rightarrow 2\ PbO(g) + 2\ SO_2(g)$

$$\text{Moles PbS} = (3.75\text{ kg PbS})\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol PbS}}{239.3\text{ g PbS}}\right) = 15.6707\text{ mol PbS (unrounded)}$$

$$\text{Moles }O_2 = n = PV / RT = \frac{(2.0\text{ atm})(228\text{ L})}{\left(0.0821\dfrac{L\cdot atm}{mol\cdot K}\right)((273+220)K)} = 11.266\text{ mol }O_2\text{ (unrounded)}$$

Dividing each mole by the coefficient in the equation [(15.6707 / 2) and (11.266 / 3)] shows that O_2 has the lower value. Thus, O_2 is limiting.

Volume SO_2 = (11.266 mol O_2) (2 mol SO_2 / 3 mol O_2) (22.4 L SO_2 / 1 mol SO_2)

    = 168.2 = **1.7 x 10^2 L SO_2**

5.61 $2 HgO(s) \rightarrow 2 Hg(l) + O_2(g)$
Moles $O_2 = n$ = amount of HgO x 1/molar mass HgO x mole ratio
Pressure $O_2 = P = nRT / V$ = (amount of HgO x 1/molar mass HgO x mole ratio)RT / V

$$P = \frac{\left[\left(40.0 \text{ g HgO}\right)\left(\frac{20.0\%}{100\%}\right)\left(\frac{1 \text{ mol HgO}}{216.6 \text{ g HgO}}\right)\left(\frac{1 \text{ mol } O_2}{2 \text{ mol HgO}}\right)\right]\left(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 25.0)\text{K}\right)}{502 \text{ mL}}\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$$

= 0.9000305 = **0.900 atm O_2**

5.62   As the temperature of the gas sample increases, the most probable speed increases. This will increase both the number of collisions per unit time and the force of each collision with the sample walls. Thus, the gas pressure increases.

5.63   At STP (or any identical temperature and pressure), the volume occupied by a mole of any gas will be identical. This is because at the same temperature, all gases have the same average kinetic energy, resulting in the same pressure.

5.64   The rate of effusion is much higher for a gas than its rate of diffusion. Effusion occurs into an evacuated space, whereas diffusion occurs into another gas. It is reasonable to expect that a gas will escape faster into a vacuum than it will into a space already occupied by another gas. The ratio of the rates of effusion and diffusion for two gases will be the same since both are inversely proportional to the square root of their molar masses.

5.65   a) Mass O_2 > mass H_2
        b) $d_{O_2} > d_{H_2}$
        c) Identical
        d) Identical
        e) Average speed H_2 > average speed O_2
        f) Effusion time H_2 < effusion time O_2

5.66   The molar masses of the three gases are 2.016 for H_2 (Flask A), 4.003 for He (Flask B), and 16.04 for CH_4 (Flask C). Since hydrogen has the smallest molar mass of the three gases, 4 g of H_2 will contain more gas molecules than 4 g of He or 4 g of CH_4. Since helium has a smaller molar mass than methane, 4 g of He will contain more gas molecules than 4 g of CH_4.
        a) **$P_A > P_B > P_C$** The pressure of a gas is proportional to the number of gas molecules. So, the gas sample with more gas molecules will have a greater pressure.
        b) **$E_A = E_B = E_C$** Average kinetic energy depends only on temperature. The temperature of each gas sample is 273 K, so they all have the same average kinetic energy.
        c) **$rate_A > rate_B > rate_C$** When comparing the speed of two gas molecules, the one with the lower mass travels faster.
        d) **total E_A > total E_B > total E_C** Since the average kinetic energy for each gas is the same (part b of this problem) then the total kinetic energy would equal the average times the number of molecules. Since the hydrogen flask contains the most molecules, its total kinetic energy will be the greatest.
        e) **$d_A = d_B = d_C$** Under the conditions stated in this problem, each sample has the same volume, 5 L, and the same mass, 4 g. Thus, the density of each is 4 g/5 L = 0.8 g/L.
        f) **Collision frequency (A) > collision frequency (B) > collision frequency (C)** The number of collisions depends on both the speed and the distance between gas molecules. Since hydrogen is the lightest molecule it has the greatest speed and the 5 L flask of hydrogen also contains the most molecules, so collisions will occur more frequently between hydrogen molecules than between helium molecules. By the same reasoning, collisions will occur more frequently between helium molecules than between methane molecules.

5.67   To find the ratio of effusion rates, calculate the inverse of the ratio of the square roots of the molar masses (equation 5.14).

$$\frac{\text{Rate } H_2}{\text{Rate } UF_6} = \sqrt{\frac{\text{Molar Mass } UF_6}{\text{Molar Mass } H_2}} = \sqrt{\frac{352.0 \text{ g/mol}}{2.016 \text{ g/mol}}} = 13.2137 = \mathbf{13.21}$$

5.68   To find the ratio of effusion rates, calculate the inverse of the ratio of the square roots of the molar masses (equation 5.14).

$$\frac{\text{Rate } O_2}{\text{Rate } Kr} = \sqrt{\frac{\text{Molar Mass } Kr}{\text{Molar Mass } O_2}} = \sqrt{\frac{83.80 \text{ g/mol}}{32.00 \text{ g/mol}}} = 1.618255 = \mathbf{1.618}$$

5.69   a) The gases have the same average kinetic energy because they are at the same temperature. The heavier Ar atoms are moving slower than the lighter He atoms to maintain the same average kinetic energy. Therefore, **Curve 1** better represents the behavior of Ar.
b) A gas that has a slower molecular speed would effuse more slowly, so **Curve 1** is the better choice.
c) Fluorine gas exists as a diatomic molecule, F_2, with $\mathcal{M}$ = 38.00 g/mol. Therefore, F_2 is much closer in size to Ar than He, so **Curve 1** more closely represents F_2's behavior.

5.70   a) **Curve 1**
b) **Curve 2**
c) **Curve 2**

5.71 $$\frac{\text{Rate He}}{\text{Rate } F_2} = \frac{\text{Time } F_2}{\text{Time He}} = \sqrt{\frac{\text{Molar Mass } F_2}{\text{Molar Mass He}}} = \sqrt{\frac{38.00 \text{ g/mol}}{4.003 \text{ g/mol}}} = 3.08105 \text{ (unrounded)}$$

Time F_2 = 3.08105 (time He) = 3.08105 (4.55 min) = 14.018786 = **14.0 min**

5.72 $$\frac{\text{Rate } H_2}{\text{Rate Unknown}} = \frac{\text{Time Unk}}{\text{Time } H_2} = \sqrt{\frac{\text{Molar Mass Unk}}{\text{Molar Mass } H_2}}$$

$$\frac{11.1 \text{ min}}{2.42 \text{ min}} = \sqrt{\frac{\text{Molar Mass Unk}}{2.016 \text{ g/mol}}}$$

$$4.586777 = \sqrt{\frac{\text{Molar Mass Unk}}{2.016 \text{ g/mol}}}$$

Molar Mass unknown = 42.41366 = **42.4 g / mol**

5.73   White phosphorus is a molecular form of the element phosphorus consisting of some number, x, of phosphorus atoms. Determine the number of phosphorus atoms, x, in one molecule of white phosphorus from the rate of effusion of the gaseous phosphorus molecules.

$$\frac{\text{Rate } P_x}{\text{Rate Ne}} = \sqrt{\frac{\text{Molar Mass Ne}}{\text{Molar Mass } P_x}} = 0.404$$

$$\frac{\text{Molar Mass Ne}}{\text{Molar Mass } P_x} = (0.404)^2 = 0.163216 \text{ (unrounded)}$$

Molar Mass P_x = Molar Mass Ne / 0.163216 = 20.18 / 0.163216 = 123.6398 g/mol
Molar Mass P_x = x (Molar Mass P) = x (30.97 g/mol)
x = (123.6398 g/mol) / (30.97 g/mol) = 3.9922 = 4
Thus, **4 atoms per molecule**, so $P_x = P_4$.

5.74   a) 0°C = 273 K   30°C = 303 K   (4.003 g He/mol) (1 kg / 10^3 g) = 0.004003 kg/mol

$$u_{rms} \text{ He (at 0°C)} = \sqrt{\dfrac{3\left(8.314\dfrac{J}{mol \bullet K}\right)(273 \text{ K})}{0.004003 \text{ kg/mol}}\left(\dfrac{kg \bullet m^2/s^2}{J}\right)} = 1.3042 \times 10^3 = \mathbf{1.30 \times 10^3 \ m/s}$$

$$u_{rms} \text{ He (at 30°C)} = \sqrt{\dfrac{3\left(8.314\dfrac{J}{mol \bullet K}\right)(303 \text{ K})}{0.004003 \text{ kg/mol}}\left(\dfrac{kg \bullet m^2/s^2}{J}\right)} = 1.3740 \times 10^3 = \mathbf{1.37 \times 10^3 \ m/s}$$

b) (131.3 g Xe/mol) (1 kg / 10^3 g) = 0.1313 kg/mol

$$u_{rms} \text{ He (at 0°C)} = \sqrt{\dfrac{3\left(8.314\dfrac{J}{mol \bullet K}\right)(303 \text{ K})}{0.1313 \text{ kg/mol}}\left(\dfrac{kg \bullet m^2/s^2}{J}\right)} = 239.913 \text{ m/s (unrounded)}$$

Rate He / Rate Xe = (1.3740 x 10^3 m/s) / (239.913 m/s) = 5.727076 = **5.72**
He molecules travel at almost 6x the speed of Xe molecules.
c) E_{He} = 1/2 mv^2 = (1/2) (0.004003 kg/mol) (1.3740 x 10^3 m/s)^2(1J/(kg•m^2/s^2) = 3778.58 = **3.78 x 10^3 J/mol**
E_{Xe} = 1/2 mv^2 = (1/2) (0.1313 kg/mol) (239.913 m/s)^2(1J/(kg•m^2/s^2) = 3778.69 = **3.78 x 10^3 J/mol**
d) (3778.58 J/mol) (1 mol He / 6.022 x 10^{23} atoms He) = 6.2746 x 10^{-21} = **6.27 x 10^{-21} J/He atom**

5.75   Intermolecular attractions cause the real pressure to be *less than* ideal pressure, so it causes a *negative* deviation. The size of the intermolecular attraction is related to the constant *a*. According to Table 5.5, a_{N_2} = 1.39, a_{Kr} = 2.32 and a_{CO_2} = 3.59. Therefore, CO_2 experiences a greater negative deviation in pressure than the other two gases: N_2 < Kr < CO_2.

5.76   Molecular size causes a positive deviation from ideal behavior. Thus, $V_{Real \ Gases}$ > $V_{Ideal \ Gases}$. Therefore, the order is H_2 < O_2 < Cl_2.

5.77   Nitrogen gas behaves more ideally at 1 atm than at 500 atm because at lower pressures the gas molecules are farther apart. An ideal gas is defined as consisting of gas molecules that act independently of the other gas molecules. When gas molecules are far apart, they act ideally, because intermolecular attractions are less important and the volume of the molecules is a smaller fraction of the container volume.

5.78   At 150°C. At higher temperatures, intermolecular attractions become less important and the volume occupied by the molecules becomes less important.

5.79   Do not forget to multiple the area of each side by two.
Surface area of can = 2 (40.0 cm) (15.0 cm) + 2 (40.0 cm) (12.5 cm) + 2 (15.0 cm) (12.5 cm)
        = 2.575 x 10^3 cm^2 (unrounded)
Total force = (2.575 x 10^3 cm^2) (1 in/2.54 cm)^2(14.7 lb/in^2) = 5.8671 x 10^3 = **5.87 x 10^3 lbs**

5.80   Molar mass has units of g/mol, so solve for the number of hemoglobin (Hb) moles combined with O_2. Moles of oxygen (from ideal gas equation) combine with Hb in a 4:1 ratio.
PV = nRT

$$\text{Moles } O_2 = PV / RT = \dfrac{(743 \text{ torr})(1.53 \text{ mL})}{\left(0.0821\dfrac{L \bullet atm}{mol \bullet K}\right)((273 + 37)\text{K})}\left(\dfrac{1 \text{ atm}}{760 \text{ torr}}\right)\left(\dfrac{10^{-3} \text{ L}}{1 \text{ mL}}\right)$$

        = 5.87708 x 10^{-5} mol O_2 (unrounded)
Moles Hb = (5.87708 x 10^{-5} mol O_2) (1 mol Hb / 4 mol O_2) = 1.46927 x 10^{-5} mol Hb (unrounded)
Molar mass hemoglobin = (1.00 g Hb) / (1.46927 x 10^{-5} mol Hb) = 6.806098 x 10^4 = **6.81 x 10^4 g/mol**

5.81   Reaction 1: 2 $NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$
Moles CO_2 = (1.00 g $NaHCO_3$) (1 mol $NaHCO_3$ / 84.01 g $NaHCO_3$) (1 mol CO_2 /2 mol $NaHCO_3$)
= 5.95167 x 10^{-3} mol CO_2 (unrounded)

$$V = nRT / P = \frac{\left(5.95167 \times 10^{-3}\, mol\right)\left(0.0821\frac{L \bullet atm}{mol \bullet K}\right)\left((273 + 200.)K\right)}{0.975\ atm}\left(\frac{1\ mL}{10^{-3}\ L}\right)$$

= 237.049 = **237 mL CO_2** Reaction 1
Reaction 2: $NaHCO_3(s) + H^+(aq) \rightarrow H_2O(l) + CO_2(g) + Na^+(aq)$
Moles CO_2 = (1.00 g $NaHCO_3$) (1 mol $NaHCO_3$ / 84.01 g $NaHCO_3$) (1 mol CO_2 /1 mol $NaHCO_3$)
= 1.1903 x 10^{-2} mol CO_2 (unrounded)

$$V = nRT / P = \frac{\left(1.1903 \times 10^{-2}\, mol\right)\left(0.0821\frac{L \bullet atm}{mol \bullet K}\right)\left((273 + 200.)K\right)}{0.975\ atm}\left(\frac{1\ mL}{10^{-3}\ L}\right)$$

= 474.0986 = **474 mL CO_2** Reaction 2

5.82   Rearrange PV = nRT to R = PV / nT
This gives: $P_iV_i / n_iT_i = P_fV_f / n_fT_f$
P_i = 1.01 atm P_f = 1.01 atm   (Thus, P has no effect, and does not need to be included.)
T_i = 305 K T_f = 250 K
$n_i = n_i$ $n_f = 0.75\ n_i$
V_i = 600. L V_f = ?

$$V_f = (n_fT_fV_i) / (n_iT_i) = \frac{\left(0.75\,n_i\left(250\ K\right)\left(600.\ L\right)\right)}{\left(n_i\right)\left(305\ K\right)} = 368.852 = \textbf{369 L}$$

5.83   a) Ideal gas equation: PV = nRT

$$P_{IGL} = nRT/V = \frac{\left[\left(0.5850\ kg\ Cl_2\right)\left(\frac{10^3\ g}{1\ kg}\right)\left(\frac{1\ mol\ Cl_2}{70.90\ g\ Cl_2}\right)\right]\left(0.08206\frac{L \bullet atm}{mol \bullet K}\right)\left((273 + 225)K\right)}{15.00\ L}$$

= 22.4791 = **22.5 atm**

b) $\left(P + \dfrac{n^2a}{V^2}\right)(V - nb) = nRT$

$P_{VDW} = \dfrac{nRT}{V - nb} - \dfrac{n^2a}{V^2}$

n = (0.5850 kg) (10^3 g/1 kg) (1 mol Cl_2 / 70.90 g Cl_2) = 8.251078 mol Cl_2 (unrounded)

$$P_{VDW} = \frac{\left(8.251078\ mol\ Cl_2\right)\left(0.08206\frac{L \bullet atm}{mol \bullet K}\right)\left((273 + 225)K\right)}{15.00\ L - \left(8.251078\ mol\ Cl_2\right)\left(0.0562\frac{L}{mol}\right)} - \frac{\left(8.251078\ mol\ Cl_2\right)^2\left(6.49\frac{atm \bullet L^2}{mol^2}\right)}{\left(15.00\ L\right)^2}$$

= 21.23252 = **21.2 atm**

5.84   a) Molar Mass I = mRT / PV =

$$\frac{\left(0.1000\text{ g}\right)\left(0.08206\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273.15 + 70.00)\text{K}\right)}{\left(0.05951\text{ atm}\right)\left(750.0\text{ mL}\right)}\left(\frac{1\text{ mL}}{10^{-3}\text{ L}}\right)$$

= 63.0905 = **63.09 g I / mol**

Molar Mass II = mRT / PV =

$$\frac{\left(0.1000\text{ g}\right)\left(0.08206\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273.15 + 70.00)\text{K}\right)}{\left(0.07045\text{ atm}\right)\left(750.0\text{ mL}\right)}\left(\frac{1\text{ mL}}{10^{-3}\text{ L}}\right)$$

= 53.293 = **53.29 g II / mol**

Molar Mass III = mRT / PV =

$$\frac{\left(0.1000\text{ g}\right)\left(0.08206\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273.15 + 70.00)\text{K}\right)}{\left(0.05767\text{ atm}\right)\left(750.0\text{ mL}\right)}\left(\frac{1\text{ mL}}{10^{-3}\text{ L}}\right)$$

= 65.10349 = **65.10 g III / mol**

b)% H in I = 100% - 85.63% = 14.37% H

% H in II = 100% - 81.10% = 18.90% H

% H in III = 100% - 82.98% = 17.02% H

Assume 100 grams of each so the mass percentages are also the grams of the element.

I

(85.63 g B) (1 mol B / 10.81 g B) = 7.921369 mol B (unrounded)

(14.37 g H) (1 mol H / 1.008 g H) = 14.25595 mol H (unrounded)

Dividing by the smaller value (7.921369) gives B = 1 and H = 1.7997

The hydrogen value is not close enough to a whole number to round. Thus, both amounts need to be multiplied by the smallest value to get near whole numbers. This value is 5. Multiplying by 5 gives B = 5 and H = 9. The empirical formula is B_5H_9, which has a formula mass of 63.12 g/mol. The empirical formula mass is near the molecular mass from part a. Therefore, the empirical and molecular formulas are both **B_5H_9**.

II

(81.10 g B) (1 mol B / 10.81 g B) = 7.50231 mol B (unrounded)

(18.90 g H) (1 mol H / 1.008 g H) = 18.750 mol H (unrounded)

Dividing by the smaller value (7.50231) gives B = 1 and H = 2.4992

The hydrogen value is not close enough to a whole number to round. Thus, both amounts need to be multiplied by the smallest value to get near whole numbers. This value is 2. Multiplying by 2 gives B = 2 and H = 5. The empirical formula is B_2H_5, which has a formula mass of 26.66 g/mol. Dividing the molecular formula mass by the empirical formula mass gives the relationship between the formulas: (53.29) / (26.66) = 2. The molecular formula is two times the empirical formula, or **B_4H_{10}**.

III

(82.98 g B) (1 mol B / 10.81 g B) = 7.6762 mol B (unrounded)

(17.02 g H) (1 mol H / 1.008 g H) = 16.8849 mol H (unrounded)

Dividing by the smaller value (7.6762) gives B = 1 and H = 2.2

The hydrogen value is not close enough to a whole number to round. Thus, both amounts need to be multiplied by the smallest value to get near whole numbers. This value is 5. Multiplying by 5 gives B = 5 and H = 11. The empirical formula is B_5H_{11}, which has a formula mass of 65.14 g/mol. The empirical formula mass is near the molecular mass from part a. Therefore, the empirical and molecular formulas are both **B_5H_{11}**.

c) $$\frac{\text{Rate } SO_2}{\text{Rate IV}} = \sqrt{\frac{\text{Molar Mass IV}}{\text{Molar Mass } SO_2}}$$

$$\frac{\left(\dfrac{250.0 \text{ mL}}{13.04 \text{ min}}\right)}{\left(\dfrac{350.0 \text{ mL}}{12.00 \text{ min}}\right)} = \sqrt{\frac{\text{Molar Mass Unk}}{64.07 \text{ g/mol}}} = 0.657318 \text{ (unrounded)}$$

Molar Mass IV = 27.6825 = **27.68 g / mol**

IV

% H in IV = 100% - 78.14% = 21.86% H

Assume 100 grams of each so the mass percentages are also the grams of the element.

(78.14 g B) (1 mol B / 10.81 g B) = 7.22849 mol B (unrounded)

(21.86 g H) (1 mol H / 1.008 g H) = 21.6865 mol H (unrounded)

Dividing by the smaller value (7.22849) gives B = 1 and H = 3.000

The empirical formula is BH_3, which has a formula mass of 13.83 g/mol. Dividing the molecular formula mass by the empirical formula mass gives the relationship between the formulas:

(27.68) / (13.83) = 2. The molecular formula is two times the empirical formula, or **B_2H_6**.

5.85   a) $V_f = V_i(T_f / T_i) (P_i / P_f)$

$V_f = V_i((273+100K) / (273+200K) (2 \text{ atm} / 1 \text{ atm})$

$V_f = V_i((373K) / (473K) (2 \text{ atm} / 1 \text{ atm}) = 1.577 \, V_i$ (unrounded)   **Increase**

b) $V_f = V_i(T_f / T_i) (P_i / P_f)$

$V_f = V_i((573K) / (373K) (1 \text{ atm} / 3 \text{ atm}) = 0.51206 \, V_i$ (unrounded)   **Decrease**

c) $V_f = V_i(T_f / T_i) (P_i / P_f)$

$V_f = V_i((400K) / (200K) (3 \text{ atm} / 6 \text{ atm}) = V_i$ **Unchanged**

d) $V_f = V_i(T_f / T_i) (P_i / P_f)$

$V_f = V_i((423K) / (573K) (0.2 \text{ atm} / 0.4 \text{ atm}) = 0.3691 \, V_i$ (unrounded)   **Decrease**

5.86   a) Partial pressures are calculated from Dalton's Law of Partial Pressures: $P_A = X_A(P_{total})$

Convert each mole percent to a mole fraction by dividing by 100%.

$P_{Nitrogen} = X_{Nitrogen} \, P_{Total} = (0.786) (1.00 \text{ atm}) (760 \text{ torr} / 1 \text{ atm}) = 597.36 = $ **597 torr N_2**

$P_{Oxygen} = X_{Oxygen} \, P_{Total} = (0.209) (1.00 \text{ atm}) (760 \text{ torr} / 1 \text{ atm}) = 158.84 = $ **159 torr O_2**

$P_{Carbon Dioxide} = X_{Carbon Dioxide} \, P_{Total} = (0.0004) (1.00 \text{ atm}) (760 \text{ torr} / 1 \text{ atm}) = 0.304 = $ **0.3 torr CO_2**

$P_{Water} = X_{Water} \, P_{Total} = (0.0046) (1.00 \text{ atm}) (760 \text{ torr} / 1 \text{ atm}) = 3.496 = $ **3.5 torr O_2**

b) Mole fractions can be calculated by rearranging Dalton's Law of Partial Pressures:

$X_A = P_A/P_{total}$ and multiply by 100 to express mole fraction as percent.

$P_{Total} = (569 + 104 + 40 + 47) \text{ torr} = 760 \text{ torr}$

N_2:   [(569 torr) / (760 torr)] x 100% = 74.7684 = **74.9 mol% N_2**

O_2:   [(104 torr) / (760 torr)] x 100% = 13.6842 = **13.7 mol% O_2**

CO_2:   [(40 torr) / (760 torr)] x 100% = 5.263 = **5.3 mol% CO_2**

H_2O: [(47 torr) / (760 torr)] x 100% = 6.1842 = **6.2 mol% CO_2**

c) Number of molecules of O_2 can be calculated using the Ideal Gas Equation and Avogadro's number.

$$\text{Molecules } O_2 = \frac{(104 \text{ torr})(0.50 \text{ L})}{\left(0.0821 \dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273+37)\text{K})} \left(\frac{1 \text{ atm}}{760 \text{ torr}}\right)\left(\frac{6.022 \times 10^{23} \, O_2 \text{ molecules}}{1 \text{ mol } O_2}\right)$$

$= 1.61892 \times 10^{21} = $ **1.6×10^{21} molecules O_2**

5.87   Atoms Rn $= (1.0 \text{ g Ra})\left(\dfrac{1 \text{ mol Ra}}{226 \text{ g Ra}}\right)\left(\dfrac{6.022 \times 10^{23} \text{ Ra atoms}}{1 \text{ mol Ra}}\right)\left(\dfrac{\left(\dfrac{1.373 \times 10^{4} \text{ Rn atoms}}{1.0 \times 10^{15} \text{ Ra atoms}}\right)}{\text{s}}\right)\left(\dfrac{3600 \text{ s}}{\text{h}}\right)\left(\dfrac{24 \text{ h}}{\text{day}}\right)$

$= 3.16094 \times 10^{15}$ Rn atoms / day (unrounded)

$V = nRT / P = \dfrac{\left[\left(3.16094 \times 10^{15} \text{ Rn atoms}\right)\left(\dfrac{1 \text{ mol Rn}}{6.022 \times 10^{23} \text{ Rn atoms}}\right)\right]\left(0.0821 \dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(273 \text{ K})}{(1.00 \text{ atm})}$

$= 1.17647 \times 10^{-7} = \textbf{1.2} \times \textbf{10}^{-7}$ **L Rn**

5.88   a) $V_f = P_i V_i T_f / T_i P_f = \dfrac{(1400. \text{ mmHg})(208 \text{ mL})(298 \text{ K})}{(286 \text{ K})(1 \text{ atm})}\left(\dfrac{1 \text{ atm}}{760 \text{ mmHg}}\right) = 399.23 = \textbf{4} \times \textbf{10}^2$ **mL**

b) $n = PV / RT = \dfrac{(1400. \text{ mmHg})(208 \text{ mL})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(286 \text{ K})}\left(\dfrac{77\% \text{ N}_2}{100\%}\right)\left(\dfrac{1 \text{ atm}}{760 \text{ mmHg}}\right)\left(\dfrac{10^{-3} \text{ L}}{1 \text{ mL}}\right)$

$= 0.0125649 = \textbf{0.013 mol N}_2$

5.89   a) Moles $= (5.14 \times 10^{15} \text{ t})(1000 \text{ kg / t})(10^3 \text{ g / kg})(1 \text{ mol / 28.8 g}) = 1.78472 \times 10^{20} = \textbf{1.78} \times \textbf{10}^{20}$ **mol gas**

b) $V = nRT / P = \dfrac{\left(1.78472 \times 10^{20} \text{ mol}\right)\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 25)\text{K})}{(1 \text{ atm})} = 4.36646 \times 10^{21} = \textbf{4} \times \textbf{10}^{21}$ **L**

c) Height $= \dfrac{\left(4.36646 \times 10^{21} \text{ L}\right)}{\left(5.100 \times 10^{8} \text{ km}^2\right)}\left(\dfrac{10^{-3} \text{ m}^3}{1 \text{ L}}\right)\left(\dfrac{1 \text{ km}}{10^3 \text{ m}}\right)^3 = 8.561686 = \textbf{9 km}$

The pressure exerted by the atmosphere is 1 atm only at the surface of the earth. Above the earth's surface, the pressure is lower than this value, so the air expands and extends much further up than 9 km.

5.90   The balanced equation and reactant amounts are given, so the first step is to identify the limiting reactant.

Moles NO_2 from Cu $= \left(4.95 \text{ cm}^3\right)\left(\dfrac{8.95 \text{ g Cu}}{\text{cm}^3}\right)\left(\dfrac{1 \text{ mol Cu}}{63.55 \text{ g Cu}}\right)\left(\dfrac{2 \text{ mol NO}_2}{1 \text{ mol Cu}}\right) = 1.394256$ mol NO_2 (unrounded)

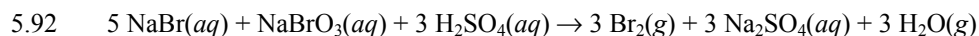Moles NO_2 from $HNO_3 = (230.0 \text{ mL})\left(\dfrac{68.0\% \text{ HNO}_3}{100\%}\right)\left(\dfrac{1 \text{ cm}^3}{1 \text{ mL}}\right)\left(\dfrac{1.42 \text{ g}}{\text{mL}}\right)\left(\dfrac{1 \text{ mol HNO}_3}{63.02 \text{ g}}\right)\left(\dfrac{2 \text{ mol NO}_2}{4 \text{ mol HNO}_3}\right)$

$= 1.7620$ mol NO_2 (unrounded)
Since less product can be made from the copper, it is the limiting reactant and excess nitric acid will be left after the reaction goes to completion. Use the calculated number of moles of NO_2 and the given temperature and pressure in the ideal gas equation to find the volume of nitrogen dioxide produced. Note that nitrogen dioxide is the only gas involved in the reaction.

$V = nRT / P = \dfrac{\left(1.394256 \text{ mol NO}_2\right)\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273.2 + 28.2)\text{K})}{(735 \text{ torr})}\left(\dfrac{760 \text{ torr}}{1 \text{ atm}}\right)$

$= 35.67427 = \textbf{35.7 L NO}_2$

5.91   a) $n = PV / RT = \dfrac{(1.0\ \text{atm})(1200\ \text{mL})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 37)\text{K}\right)}\left(\dfrac{10^{-3}\ \text{L}}{1\ \text{mL}}\right) = 0.047149 = \textbf{0.047 mol air}$

b) Molecules = (0.047149 mol air) (6.022 x 10^{23} molecules air / mol air)
   = 2.83931 x 10^{22} = **2.8 x 10^{22} molecules air**

5.92 $5\ \text{NaBr}(aq) + \text{NaBrO}_3(aq) + 3\ \text{H}_2\text{SO}_4(aq) \rightarrow 3\ \text{Br}_2(g) + 3\ \text{Na}_2\text{SO}_4(aq) + 3\ \text{H}_2\text{O}(g)$

Moles Br_2 from NaBr $= (275\ \text{g NaBr})\left(\dfrac{1\ \text{mol NaBr}}{102.89\ \text{g NaBr}}\right)\left(\dfrac{3\ \text{mol Br}_2}{5\ \text{mol NaBr}}\right) = 1.60365\ \text{mol Br}_2$ (unrounded)

Moles Br_2 from $\text{NaBrO}_3 = (175.6\ \text{g NaBrO}_3)\left(\dfrac{1\ \text{mol NaBrO}_3}{180.89\ \text{g NaBrO}_3}\right)\left(\dfrac{3\ \text{mol Br}_2}{1\ \text{mol NaBrO}_3}\right)$

   $= 2.912267\ \text{mol Br}_2$ (unrounded)
The NaBr is limiting.

Volume $= nRT / P = \dfrac{(1.60365\ \text{mol})\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 300)\text{K}\right)}{(0.855\ \text{atm})} = 88.235 = \textbf{88.2 L Br}_2$
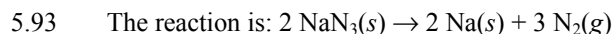
5.93   The reaction is: $2\ \text{NaN}_3(s) \rightarrow 2\ \text{Na}(s) + 3\ \text{N}_2(g)$
The vapor pressure of water (Table 5.3) at 26°C is 25.2 torr (= 25.2 mmHg).
Volume = nRT/ P =

$\dfrac{\left[50.0\ \text{g NaN}_3\left(\dfrac{1\ \text{mol NaN}_3}{65.02\ \text{g NaN}_3}\right)\left(\dfrac{3\ \text{mol N}_2}{2\ \text{mol NaN}_3}\right)\right]\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 26)\text{K}\right)}{\left((745.5 - 25.2)\text{mmHg}\right)}\left(\dfrac{760\ \text{mmHg}}{1\ \text{atm}}\right)$

   $= 29.8764 = \textbf{29.9 L N}_2$

5.94   It is necessary to determine both the empirical formula and the molar mass of the gas.
        Empirical formula:
        Assume 100.0 grams of sample to make the percentages of each element equal to the grams of that
            element. Thus, 64.81% C = 64.81 g C, 13.60% H = 13.60 g H, and 21.59% O = 21.59 g O.
        Moles C = (64.81 g C) (1 mol C / 12.01 g C) = 5.39634 mol C (unrounded)
        Moles H = (13.60 g H) (1 mol H / 1.008 g H) = 13.49206 mol H (unrounded)
        Moles O = (21.59 g O) (1 mol O / 16.00 g O) = 1.349375 mol O (unrounded)
        Divide by the smallest number of moles (O):
        C = (5.39634 mol) / (1.349375 mol) = 4
        H = (13.49206 mol) / (1.349375 mol) = 10
        O = (1.349375 mol) / (1.349375 mol) = 1
        Empirical formula = $\text{C}_4\text{H}_{10}\text{O}$ (empirical formula mass = 74.12 g/mol)
        Molar Mass:

$\mathcal{M} = mRT / PV = \dfrac{(2.57\ \text{g})\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 25)\text{K}\right)}{(0.420\ \text{atm})(2.00\ \text{L})} = 74.85\ \text{g/mol}$ (unrounded)

        Molecular formula:
        Since the molar mass and the empirical formula mass are similar, the empirical and molecular formulas
        must both be: **$\text{C}_4\text{H}_{10}\text{O}$**

5.95 The empirical formula for aluminum chloride is $AlCl_3$ ($\mathcal{M}$ = 133.33 g/mol). Calculate the molar mass of the gaseous species from the ratio of effusion rates. This molar mass, divided by the empirical weight, should give a whole number multiple that will yield the molecular formula.

$$\frac{\text{Rate Unk}}{\text{Rate He}} = 0.122 = \sqrt{\frac{\text{Molar Mass He}}{\text{Molar Mass Unk}}}$$

$$0.122 = \sqrt{\frac{4.003 \text{ g/mol}}{\text{Molar Mass Unk}}}$$

Molar Mass Unknown = 270.7605 g/mol (unrounded)
The whole number multiple is 270.7605/133.33, which is about 2. Therefore, the molecular formula of the gaseous species is 2 x $(AlCl_3)$ = **Al_2Cl_6**.

5.96 a) $2 C_8H_{18}(l) + 25 O_2(g) \rightarrow 16 CO_2(g) + 18 H_2O (g)$

$$\text{Moles products} = \left(100.\text{ g } C_8H_{18}\right)\left(\frac{1 \text{ mol } C_8H_{18}}{114.22 \text{ g } C_8H_{18}}\right)\left(\frac{34 \text{ mol gas}}{2 \text{ mol } C_8H_{18}}\right) = 14.883558 \text{ mol gas (unrounded)}$$

$$V = nRT / P = \frac{\left(14.883558 \text{ mol gas}\right)\left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 350)\text{K}\right)}{\left(735 \text{ torr}\right)}\left(\frac{760 \text{ torr}}{1 \text{ atm}}\right)$$

$$= 787.162 = \textbf{787 L gas}$$

b) $\text{Moles } O_2 = \left(100.\text{ g } C_8H_{18}\right)\left(\frac{1 \text{ mol } C_8H_{18}}{114.22 \text{ g } C_8H_{18}}\right)\left(\frac{25 \text{ mol } O_2}{2 \text{ mol } C_8H_{18}}\right) = 10.94379 \text{ mol } O_2 \text{ (unrounded)}$

Moles other gases = (10.94379 mol O_2) [(78 + 1.0)%] / (21%) = 41.1695 mol gas (unrounded)

$$V = nRT / P = \frac{\left(41.1695 \text{ mol gas}\right)\left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 350)\text{K}\right)}{\left(735 \text{ torr}\right)}\left(\frac{760 \text{ torr}}{1 \text{ atm}}\right)$$

$$= 2177.37 \text{ L residual air (unrounded)}$$

Total volume of gaseous exhaust = 787.162 L + 2177.37 L = 2964.53 = **2.96 x 10³ L**

5.97 Plan: First, write the balanced equation for the reaction: $2 SO_2 + O_2 \rightarrow 2 SO_3$. The total number of moles of gas will change as the reaction occurs since 3 moles of reactant gas forms 2 moles of product gas. From the volume, temperature and pressures given, we can calculate the number of moles of gas before and after the reaction. For each mole of SO_3 formed, the total number of moles of gas decreases by 1/2 mole. Thus, twice the decrease in moles of gas equals the moles of SO_3 formed.
Solution:
Moles of gas before and after reaction

$$\text{Initial moles} = PV / RT = \frac{\left(1.95 \text{ atm}\right)\left(2.00 \text{ L}\right)}{\left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left(900.\text{ K}\right)} = 0.05278116 \text{ mol (unrounded)}$$

$$\text{Final moles} = PV / RT = \frac{\left(1.65 \text{ atm}\right)\left(2.00 \text{ L}\right)}{\left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left(900.\text{ K}\right)} = 0.04466098 \text{ mol (unrounded)}$$

Moles of SO_3 produced = 2 x decrease in the total number of moles
    = 2 x (0.05278116 mol - 0.04466098 mol)
    = 0.01624035 = **1.62 x 10⁻² mol**
Check: If the starting amount is 0.0528 total moles of SO_2 and O_2, then x + y = 0.0528 mol, where x = mol of SO_2 and y = mol of O_2. After the reaction:
(x - z) + (y - 0.5z) + z = 0.0447 mol
Where z = mol of SO_3 formed = mol of SO_2 reacted = 2(mol of O_2 reacted).

Subtracting the two equations gives:

$$x - (x - z) + y - (y - 0.5z) - z = 0.0528 - 0.0447$$
$$z = 0.0163 \text{ mol } SO_3$$

The approach of setting up two equations and solving them gives the same result as above.

5.98   The reaction is: $2\ NCl_3(l) \xrightarrow{\Delta} N_2(g) + 3\ Cl_2(g)$

The decomposition of all the NCl_3 means that the final pressure must be due to the N_2 and the Cl_2. The product gases are present at a 1:3 ratio, and the total moles are 4.

a) Partial pressure $N_2 = P_{nitrogen} = X_{nitrogen}\ P_{total} = (1\ \text{mol } N_2\ /\ 4\ \text{mol total})\ (754\ \text{mmHg})$
$$= 188.5 = \textbf{188 mmHg } \mathbf{N_2}$$

Partial pressure $Cl_2 = P_{nitrogen} = X_{nitrogen}\ P_{total} = (3\ \text{mol } Cl_2\ /\ 4\ \text{mol total})\ (754\ \text{mmHg}) = 565.5 = \textbf{566 mmHg } \mathbf{Cl_2}$

b) The mass of NCl_3 may be determined several ways. Using the partial pressure of Cl_2 gives:

$$PV = mRT\ /\ \mathcal{M}$$

$$m = PV\ \mathcal{M}\ /\ RT = \frac{(565.5\ \text{mmHg } Cl_2)(2.50\ \text{L})(120.36\ \text{g/mol})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 95)\text{K})}\left(\frac{1\ \text{atm}}{760\ \text{mmHg}}\right)\left(\frac{2\ \text{mol } NCl_3}{3\ \text{mol } Cl_2}\right)$$

$$= 4.94036 = \textbf{4.94 g } \mathbf{NCl_3}$$

5.99   The reaction is: $2\ NH_4NO_3(s) \rightarrow 2\ N_2(g) + O_2(g) + 4\ H_2O(g)$
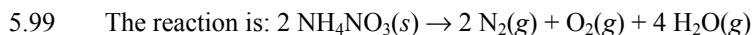
$$\text{Moles gas} = (15.0\ \text{kg } NH_4NO_3)\left(\frac{10^3\ \text{g}}{1\ \text{kg}}\right)\left(\frac{1\ \text{mol } NH_4NO_3}{80.05\ \text{g } NH_4NO_3}\right)\left(\frac{7\ \text{mol Gas}}{2\ \text{mol } NH_4NO_3}\right)$$

$$= 655.840\ \text{mol gas (unrounded)}$$

$$V = nRT\ /\ P = \frac{(655.840\ \text{mol Gas})\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 307)\text{K})}{(1.00\ \text{atm})} = 3.1229789 \times 10^4 = \textbf{3.12 x } \mathbf{10^4}\ \textbf{L}$$

5.100   The balanced equation is: $I_2(aq) + 2\ S_2O_3^{2-}(aq) \rightarrow 2\ I^-(aq) + S_4O_6^{2-}(aq)$
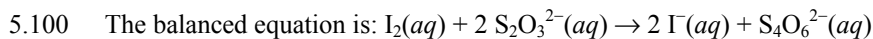
$$\text{Initial moles of } I_2 = (20.00\ \text{mL})\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)\left(\frac{0.01017\ \text{mol } I_2}{\text{L}}\right) = 2.034 \times 10^{-4}\ \text{mol } I_2\ \text{initial}$$

$$\text{Moles } I_2 \text{ reacting with } S_2O_3^{2-} = (11.37\ \text{mL})\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)\left(\frac{0.0105\ \text{mol } S_2O_3^{2-}}{\text{L}}\right)\left(\frac{1\ \text{mol } I_2}{2\ \text{mol } S_2O_3^{2-}}\right)$$

$$= 5.96925 \times 10^{-5}\ \text{mol } I_2 \text{ reacting with } S_2O_3^{2-} \text{ (not reacting with } SO_2)$$

Moles I_2 reacting with $SO_2 = 2.034 \times 10^{-4}\ \text{mol} - 5.96925 \times 10^{-5}\ \text{mol}$
$$= 1.437075 \times 10^{-4}\ \text{mol } I_2 \text{ reacted with } SO_2 \text{ (unrounded)}$$

The balanced equation is: $SO_2(aq) + I_2(aq) + 2\ H_2O(l) \rightarrow HSO_4^-(aq) + 2\ I^-(aq) + 3\ H^+(aq)$

Moles $SO_2 = (1.437075 \times 10^{-4}\ \text{mol } I_2)\ (1\ \text{mol } SO_2\ /\ 1\ \text{mol } I_2) = 1.437075 \times 10^{-4}\ \text{mol } SO_2 \text{ (unrounded)}$

$$\text{Moles air} = \frac{(700.\ \text{torr})(500.\ \text{mL})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 38)\text{K})}\left(\frac{1\ \text{atm}}{760\ \text{torr}}\right)\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right) = 0.018036\ \text{mol air (unrounded)}$$

Volume % $SO_2 = \text{mol \% } SO_2 = [(1.437075 \times 10^{-4}\ \text{mol } SO_2)\ /\ (0.018036\ \text{mol air})] \times 100\%$
$$= 0.796781 = \textbf{0.797\%}$$

5.101 a) The balanced equation for the reaction is $Ni(s) + 4\ CO(g) \rightarrow Ni(CO)_4(g)$. Although the problem states that Ni is impure, you can assume the impurity is unreactive and thus not include it in the reaction. The mass of Ni can be calculated from the stoichiometric relationship between moles of CO (using the ideal gas equation) and Ni.

Mass Ni = mol Ni x molar mass

$$\text{Mass Ni} = \frac{(100.7\ \text{kPa})(3.55\ \text{m}^3)}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\big((273 + 50)\,\text{K}\big)}\left(\frac{1\ \text{atm}}{101.325\ \text{kPa}}\right)\left(\frac{1\ \text{L}}{10^{-3}\ \text{m}^3}\right)\left(\frac{1\ \text{mol Ni}}{4\ \text{mol CO}}\right)\left(\frac{58.69\,\text{g Ni}}{1\,\text{mol Ni}}\right)$$

$= 1952.089 =$ **1.95 x 10³ g Ni**

b) Assume the volume is 1 m³. Use the ideal gas equation to solve for moles of $Ni(CO)_4$, which equals the moles of Ni, and convert moles to grams using the molar mass.

$$\text{Mass Ni} = \frac{(21\ \text{atm})(\text{m}^3)}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\big((273 + 155)\,\text{K}\big)}\left(\frac{1\ \text{L}}{10^{-3}\ \text{m}^3}\right)\left(\frac{1\ \text{mol Ni}}{1\ \text{mol Ni(CO)}_4}\right)\left(\frac{58.69\ \text{g Ni}}{1\ \text{mol Ni}}\right)$$

$= 3.50749\text{ x }10^4 =$ **3.5 x 10⁴ g Ni**

The pressure limits the significant figures.

c) The amount of CO needed to form $Ni(CO)_4$ is the same amount that is released on decomposition. The vapor pressure of water at 35°C (42.2 torr) must be accounted for (see Table 5.3). Use the ideal gas equation to calculate the volume of CO, and compare this to the volume of $Ni(CO)_4$. The mass of Ni from a m³ (part b) can be used to calculate the amount of CO released.

$P_{CO} = P_{total} - P_{water} = 769\ \text{torr} - 42.2\ \text{torr} = 726.8\ \text{torr}$ (unrounded)

$V = nRT / P =$

$$\frac{\left[3.50749\text{ x }10^4\ \text{g Ni}\left(\dfrac{1\ \text{mol Ni}}{58.69\ \text{g Ni}}\right)\left(\dfrac{4\ \text{mol CO}}{1\ \text{mol Ni}}\right)\right]\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\big((273 + 35)\,\text{K}\big)}{726.8\ \text{torr}}\left(\frac{760\ \text{torr}}{1\ \text{atm}}\right)\left(\frac{10^{-3}\ \text{m}^3}{1\ \text{L}}\right)$$

$= 63.209866 =$ **63 m³ CO**

The answer is limited to two significant figures because the mass of Ni comes from part b.

5.102 Assume 100 grams of sample, thus 33.01% Si is equivalent to 33.01 grams Si, and 66.99% F is equivalent to 66.99 grams of F.

Moles Si = (33.01 g Si) (1 mol Si / 28.09 g Si) = 1.17515 mol Si (unrounded)
Moles F = (66.99 g F) (1 mol F / 19.00 g F) = 3.525789 mol F (unrounded)
Dividing by the smaller value (1.17515) gives an empirical formula of SiF_3 (empirical formula mass = 85 g/mol).

$$\mathcal{M} = \text{mRT / PV} = \frac{(2.60\ \text{g})\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\big((273 + 27)\,\text{K}\big)}{(1.50\ \text{atm})(0.250\ \text{L})} = 170.768\ \text{g/mol (unrounded)}$$

The molar mass is twice the empirical formula mass, so the molecular formula must be twice the empirical formula, or 2 x SiF_3 = **Si₂F₆**.

5.103 a) A preliminary equation for this reaction is $4\ C_xH_yN_z + n\ O_2 \rightarrow 4\ CO_2 + 2\ N_2 + 10\ H_2O$.
Since the organic compound does not contain oxygen, the only source of oxygen as a reactant is oxygen gas. To form 4 volumes of CO_2 would require 4 volumes of O_2 and to form 10 volumes of H_2O would require 5 volumes of O_2. Thus, 9 **volumes of O₂** was required.
b) Since the volume of a gas is proportional to the number of moles of the gas we can equate volume and moles. From a volume ratio of 4 CO_2:2 N_2:10 H_2O we deduce a mole ratio of 4 C:4 N:20 H or 1 C:1 N:5 H for an empirical formula of **CH₅N**.

5.104   The final pressure will be the sum of the air pressure (0.980 atm) and the pressure generated by the carbon
        dioxide.

$$\text{Carbon dioxide pressure} = nRT / V = \frac{\left[10.0 \text{ g CO}_2\left(\dfrac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2}\right)\right]\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273.2 + 550.0)\text{K}\right)}{(0.800 \text{ L})}$$

            = 19.1958 atm (unrounded)
        P_{total} = (0.980 atm) + (19.1958 atm) = 20.1758 = **20.2 atm**

5.105   a) **2 x 10⁶ blue** particles and **2 x 10⁶ black** particles
        b) **2 x 10⁶ blue** particles and **2 x 10⁶ black** particles
        c) Final pressure in C = (2/3) (750 torr) = **500 torr**
        d) Final pressure in B = (2/3) (750 torr) = **500 torr**

5.106   At this altitude, the atmosphere is very thin. The collision frequency is very low and thus there is very little transfer
        of kinetic energy. Satellites and astronauts would not become "hot" in the usual sense of the word.

5.107   To find the factor by which a diver's lungs would expand, find the factor by which P changes from 125 ft to the
        surface, and apply Boyle's Law. To find that factor, calculate $P_{seawater}$ at 125 ft by converting the given depth from
        ft-seawater to mmHg to atm and adding the surface pressure (1.00 atm).
        P (H_2O) = (125 ft) (12 in / ft) (2.54 cm / 1 in) (10^{-2} m / 1 cm) (1 mm / 10^{-3} m) = 3.81 x 10^4 mmH$_2$O
        P (Hg) = (3.81 x 10^4 mmH$_2$O) [(1.04 g H$_2$O/mL) / (13.5 g Hg /mL)] (1 atm / 760 mmHg)
                = 3.861988 atm (unrounded)
        P_{total} = (1.00 atm) + (3.861988 atm) = 4.861988 atm (unrounded)
        Use Boyle's Law to find the volume change of the diver's lungs:
        $V_f = V_i P_i / P_f = V_i$ (4.861988 atm) / (1.00 atm) = 4.861988 = **4.86 V$_i$**
        To find the depth to which the diver could ascend safely, use the given safe expansion factor (1.5) and the pressure
        at 125 ft, P_{125}, to find the safest ascended pressure, P_{safe}.
                P_{125} / P_{safe} = 1.5
                $P_{safe} = P_{125}$ / 1.5 = (4.861988 atm) / 1.5 = 3.241325 atm (unrounded)
        Convert the pressure in atm to pressure in ft of seawater using the conversion factors above. Subtract this distance
        from the initial depth to find how far the diver could ascend.
        P (H_2O) = [(4.861988 - 3.241325)atm] x (760 mmHg / 1 atm) x [(13.5 g Hg/mL) / (1.04 g H$_2$O/mL)]
                x (10^{-3} m / 1 mm) x (1.094 yd / 1 m) x (3 ft / 1 yd) = 52.4741 ft (unrounded)
        Therefore, the diver can safely ascend 52.5 ft to a depth of (125 - 52.4741) = 72.5259 = **73 ft**.

5.108   The balanced equation is: $CaF_2(s) + H_2SO_4(aq) \rightarrow 2 \text{ HF}(g) + CaSO_4(s)$

$$T = PV / nR = \frac{(875 \text{ torr})(8.63 \text{ L})}{\left[15.0 \text{ g CaF}_2\left(\dfrac{1 \text{ mol CaF}_2}{78.08 \text{ g CaF}_2}\right)\left(\dfrac{2 \text{ mol HF}}{1 \text{ mol CaF}_2}\right)\right]\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)}\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right)$$

            = 314.9783 K
        The gas must be heated to **315 K**.

5.109 First, write the balanced equation: $2 H_2O_2(aq) \rightarrow 2 H_2O(l) + O_2(g)$

According to the description in the problem, a given volume of peroxide solution (0.100 L) will release a certain number of "volumes of oxygen gas" (20). Assume that 20 is exact.

0.100 L solution will produce (20 x 0.100 L) = 2.00 L O_2 gas

You can convert volume of $O_2(g)$ to moles of gas using the ideal gas equation.

$$\text{Moles } O_2 = PV / RT = \frac{(1.00 \text{ atm})(2.00 \text{ L})}{\left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(273 \text{ K})} = 8.92327 \times 10^{-2} \text{ mol } O_2 \text{ (unrounded)}$$

The grams of hydrogen peroxide can now be solved from the stoichiometry using the balanced chemical equation.

Mass H_2O_2 = (8.92327 x 10^{-2} mol O_2) (2 mol H_2O_2 / 1 mol O_2) (34.02 g H_2O_2 / 1 mol H_2O_2)

= 6.071395 = **6.07 g H_2O_2**

5.110 The moles may be found using the ideal gas equation. Moles times Avogadro's number gives the molecules.

$$\text{Molecules} = \frac{(10^{-8} \text{ mmHg})(1 \text{ mL})}{\left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(500 \text{ K})}\left(\frac{1 \text{ atm}}{760 \text{ mmHg}}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}}\right) =$$

$$= 1.93025 \times 10^8 = 10^8 \text{ molecules (The } 10^{-8} \text{ mmHg limits the significant figures.)}$$

5.111 a) $u_{rms} = \sqrt{\frac{3(8.314 \text{ J/mol} \cdot \text{K})(273 \text{ K})}{(32.00 \text{ g/mol})}\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{\text{kg} \cdot \text{m}^2/\text{s}^2}{\text{J}}\right)} = 461.2878 = \textbf{461 m/s}$

b) Collision frequency = (u_{rms} / mean free path) = (461.2878 m/s) / (6.33 x 10^{-8} m) = 7.2873 x 10^9 = **7.29 x 10^9 s^{-1}**

5.112 Molar volume uses exactly one mole of gas along with the temperature and pressure given in the problem.

$$V / n = RT / P = \frac{\left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(730. \text{ K})}{(90 \text{ atm})} = 0.66592 = \textbf{0.67 L/mol}$$

5.113
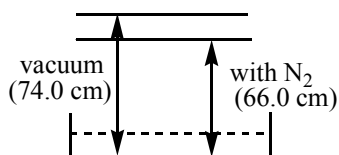
The diagram above describes the two Hg height levels within the barometer. To find the mass of N_2, find P and V (T given) and use the ideal gas equation. The P_{N_2} is directly related to the change in column height of Hg. The volume of the space occupied by the $N_2(g)$ is calculated from the dimensions of the barometer.

Pressure of the nitrogen = (74.0 - 66.0) cmHg (10^{-2} m / 1 cm) (1 mm / 10^{-3} m) (1 atm / 760 mmHg)

= 0.10526 atm (unrounded)

(Notice that the subtraction limits the results to two significant figures.)

Volume of the nitrogen = (1.00 x 10^2 - 66.0) cm (1.20 cm^2) (1 mL / 1 cm^3) (10^{-3} L / 1 mL) =

= 0.0408 L (unrounded)

The moles of nitrogen may be found using the ideal gas equations, and the moles times the molar mass of nitrogen will give the mass of nitrogen.

$$\text{Mass of nitrogen} = \frac{(0.10526 \text{ atm})(0.0408 \text{ L})}{\left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 24)\text{K})}\left(\frac{28.02 \text{ g } N_2}{1 \text{ mol}}\right)$$

= 4.935 x 10^{-3} = **4.94 x 10^{-3} g N_2**

5.114 Molar concentration is moles per liter. Use the original volume (10.0 L) to find the moles of gas, and the final volume (0.750 L) to get the molar concentration.

$$\text{Moles } NH_3 = PV / RT = \frac{(735 \text{ torr})(10.0 \text{ L})}{\left(0.0821 \dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 31)\text{K})} \left(\frac{1 \text{ atm}}{760 \text{ torr}}\right)$$

   = 0.38748688 mol (unrounded)

Molar concentration = mol / volume, in liters, of solution.

M = (0.38748688 mol) / (0.750 L) = 0.516649 = **0.517 M**

5.115 Determine the total moles of gas produced. The total moles times the fraction of each gas gives the moles of that gas which may be converted to metric tons.

$$\text{Moles total} = n = PV / RT = \frac{(1.00 \text{ atm})(1.5 \times 10^3 \text{ m}^3/\text{d})}{\left(0.0821 \dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(298 \text{ K})} \left(\frac{1 \text{ L}}{10^{-3} \text{ m}^3}\right)\left(\frac{365.25 \text{ d}}{1 \text{ y}}\right)$$

   = 2.23935×10^7 mol/year (unrounded)

Mass CO_2 = (0.4896) (2.23935×10^7 mol/year) (44.01 g CO_2/mol) (1 kg/10^3 g) (1 T/10^3 kg)
   = 482.519 = **4.8×10^2 T CO_2/y**

Mass CO = (0.0146) (2.23935×10^7 mol/year) (28.01 g CO/mol) (1 kg/10^3 g) (1 T/10^3 kg)
   = 9.15773 = **9.16×10^2 T CO/y**

Mass H_2O = (0.3710) (2.23935×10^7 mol/year) (18.02 g H_2O/mol) (1 kg/10^3 g) (1 T/10^3 kg)
   = 149.70995 = **1.50×10^2 T H_2O/y**

Mass SO_2 = (0.1185) (2.23935×10^7 mol/year) (64.07 g SO_2/mol) (1 kg/10^3 g) (1 T/10^3 kg)
   = 170.018 = **1.70×10^2 T SO_2/y**

Mass S_2 = (0.0003) (2.23935×10^7 mol/year) (64.14 g S_2/mol) (1 kg/10^3 g) (1 T/10^3 kg)
   = 0.4308957482.519 = **4×10^{-1} T S_2/y**

Mass H_2 = (0.0047) (2.23935×10^7 mol/year) (2.016 g H_2/mol) (1 kg/10^3 g) (1 T/10^3 kg)
   = 0.21218 = **2.1×10^{-1} T H_2/y**

Mass HCl = (0.0008) (2.23935×10^7 mol/year) (36.46 g HCl/mol) (1 kg/10^3 g) (1 T/10^3 kg)
   = 0.6531736 = **6×10^{-1} T CO_2/y**

Mass H_2S = (0.0003) (2.23935×10^7 mol/year) (34.09 g H_2S/mol) (1 kg/10^3 g) (1 T/10^3 kg)
   = 0.229018 = **2×10^{-1} T CO_2/y**

5.116 The balanced chemical equation is: $2 H_2(g) + O_2(g) \rightarrow 2 H_2O(l)$

Moles H_2 = (28.0 mol H_2O) (2 mol H_2/2 mol H_2O) = 28.0 mol H_2

Moles O_2 = (28.0 mol H_2O) (1 mol O_2/2 mol H_2O) = 14.0 mol O_2

a) P = nRT / V

$$P\ H_2 = \frac{(28.0 \text{ mol } H_2)\left(0.0821 \dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273.2 + 23.8)\text{K})}{(20.0 \text{ L})} = 34.137 = \textbf{34.1 atm } H_2$$

$$P\ O_2 = \frac{(14.0 \text{ mol } O_2)\left(0.0821 \dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273.2 + 23.8)\text{K})}{(20.0 \text{ L})} = 17.06859 = \textbf{17.1 atm } O_2$$

b) $\left(P + \dfrac{n^2 a}{V^2}\right)(V - nb) = nRT$

Use Table 5.5 to find the values of the van der Waals constants.

H_2: a = 0.244 atm•L^2/mol^2 b = 0.0266 L/mol

O_2: a = 1.36 atm•L^2/mol^2 b = 0.0318 L/mol

$$P_{VDW} = \frac{nRT}{V - nb} - \frac{n^2 a}{V^2}$$

$$P_{VDW}\ H_2 = \frac{(28.0\ \text{mol}\ H_2)\left(0.08206\dfrac{L \bullet atm}{mol \bullet K}\right)\big((273.2 + 23.8)K\big)}{20.0\ L - (28.0\ \text{mol}\ H_2)\left(0.0266\dfrac{L}{mol}\right)} - \frac{(28.0\ \text{mol}\ H_2)^2\left(0.244\dfrac{atm \bullet L^2}{mol^2}\right)}{(20.0\ L)^2}$$

$= 34.9631 = $ **35.0 atm H_2**

$$P_{VDW}\ O_2 = \frac{(14.0\ \text{mol}\ O_2)\left(0.08206\dfrac{L \bullet atm}{mol \bullet K}\right)\big((273.2 + 23.8)K\big)}{20.0\ L - (14.0\ \text{mol}\ O_2)\left(0.0318\dfrac{L}{mol}\right)} - \frac{(14.0\ \text{mol}\ O_2)^2\left(1.36\dfrac{atm \bullet L^2}{mol^2}\right)}{(20.0\ L)^2}$$

$= 16.78228 = $ **16.8 atm O_2**

c) The van der Waals value for hydrogen is slightly higher than the value from the ideal gas equation. The van der Waals value for oxygen is slightly lower than the value from the ideal gas equation.

5.117   a) **Xenon** would show greater deviation from ideal behavior than argon since xenon is a larger atom than xenon. The electron cloud of Xe is more easily distorted so intermolecular attractions are greater. Xe's larger size also means that the volume the gas occupies becomes a greater proportion of the container's volume at high pressures.
b) **Water** vapor would show greater deviation from ideal behavior than neon gas since the attractive forces between water molecules are greater than the attractive forces between neon atoms. We know the attractive forces are greater for water molecules because it remains a liquid at a higher temperature than neon (water is a liquid at room temperature while neon is a gas at room temperature).
c) **Mercury** vapor would show greater deviation from ideal behavior than radon gas since the attractive forces between mercury atoms is greater than that between radon atoms. We know that the attractive forces for mercury are greater because it is a liquid at room temperature while radon is a gas.
d) **Water** is a liquid at room temperature; methane is a gas at room temperature (think about where you have heard of methane gas before - Bunsen burners in lab, cows' digestive system). Therefore, water molecules have stronger attractive forces than methane molecules and should deviate from ideal behavior to a greater extent than methane molecules.

5.118   Moles of HBr in the Hydrobromic acid: (1.20 mol HBr / L) (3.50 L) = 4.20 mol HBr

$$V\ HBr = nRT / P = \frac{(4.20\ \text{mol}\ HBr)\left(0.0821\dfrac{L \bullet atm}{mol \bullet K}\right)\big((273 + 27)K\big)}{(0.975\ atm)} = 106.098 = \textbf{106 L HBr}$$

5.119   The mole fraction of CO must be found from the moles of the gases. The mole fraction plus the total pressure can be used to find the partial pressure.
Moles CO = (7.0 g CO) (1 mol CO / 28.01 g CO) = 0.24991 mol CO (unrounded)
Moles SO_2 = (10.0 g CO) (1 mol SO_2 / 64.07 g SO_2) = 0.156079 mol SO_2 (unrounded)
P CO = X_{CO} P_{total} = [(0.24991 mol CO) /(0.24991 + 0.156079)mol](0.33 atm) = 0.20313 = **0.20 atm CO**

5.120   First, balance the equation:

$$4\ FeS_2(s) + 11\ O_2(g) \rightarrow 8\ SO_2(g) + 2\ Fe_2O_3(s)$$
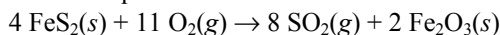
Pressure of N_2 = 1.05 atm - 0.64 atm O_2 = 0.41 **atm N_2**

Pressure of unreacted O_2 = (0.64 atm O_2) [(100 - 85)% / 100%] = **0.096 atm O_2**

Pressure of SO_2 produced = (0.64 atm O_2) (8 atm SO_2 / 11 atm O_2) = 0.46545 = **0.47 atm SO_2**

Total Pressure = (0.41 atm) + (0.096 atm) + (0.46545 atm) = 0.97145 = **0.97 atm total**

5.121   V and T are not given, so the ideal gas equation cannot be used. The total pressure of the mixture is given:

$$P_{total} = P_{krypton} + P_{carbon\ dioxide} = 0.708\ atm$$

The NaOH absorbed the CO_2 leaving the Kr, thus $P_{krypton}$ = 0.250 atm

$$P_{carbon\ dioxide} = P_{total} - P_{krypton} = 0.708\ atm - 0.250\ atm = 0.458\ atm$$

Determining mole fractions:

Carbon dioxide: X = (0.458 atm) / (0.708 atm) = 0.64689 (unrounded)

Krypton: X = (0.250 atm) / (0.708 atm) = 0.353107 (unrounded)

Relative mass fraction = [(0.353107) (83.80 g Kr/mol)] / [(0.64689) (44.01 g CO_2 / mol)]
        = 1.039366 (unrounded)

35.0 g = x g CO_2 + (1.039366 x) g Kr

Grams CO_2 = (35.0 g) / (2.039366) = 17.16219581 = **17.2 g CO_2**

Grams Kr = 35.0 g - 17.162 g CO_2 = 17.83780419 = **17.8 g Kr**

5.122   As the car accelerates, the air in the car is pressed toward the back. The helium balloon floats on "top" of this denser air, which pushes it toward the front of the car.
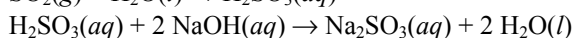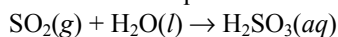
5.123   P = nRT / V =

$$\left[\left(\frac{8.0\times 10^{12}\ \text{molecules}}{\text{m}^3}\right)\left[\frac{1\ \text{mol}}{6.022\times 10^{23}\ \text{molecules}}\right]\right]\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)\left((273+22)\text{K}\right)\left(\frac{10^{-3}\ \text{m}^3}{1\ \text{L}}\right)$$
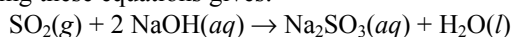
        = 3.217469 x 10^{-13} = **3.2 x 10^{-13} atm •OH**

The mole percent is the same as the pressure percent as long as the other conditions are the same.

Mole percent = [(3.217469 x 10^{-13} atm) / (1.00 atm)] 100% = 3.217469 x 10^{-11} = **3.2 x 10^{-11}%**

5.124   The balanced chemical equations are:

$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$$
$$H_2SO_3(aq) + 2\ NaOH(aq) \rightarrow Na_2SO_3(aq) + 2\ H_2O(l)$$

Combining these equations gives:

$$SO_2(g) + 2\ NaOH(aq) \rightarrow Na_2SO_3(aq) + H_2O(l)$$

M NaOH = [(mol SO_2) (2 mol NaOH / 1 mol SO_2) / (L NaOH solution)

Moles SO_2 = PV / RT

M NaOH = [(PV / RT) (2 mol NaOH / 1 mol SO_2) / (L NaOH solution)

$$M\ \text{NaOH} = \frac{\left[\frac{(740.\ \text{mmHg})(0.200\ \text{L}\ SO_2)}{\left(0.0821\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)((273+20.)\text{K})}\left(\frac{1\ \text{atm}}{760\ \text{mmHg}}\right)\right]\left(\frac{2\ \text{mol NaOH}}{1\ \text{mol}\ SO_2}\right)}{(10.0\ \text{mL})\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)}$$

        = 1.619076 = **1.62 M NaOH**

5.125   Write the balanced chemical equations:

$$2 \text{ LiH}(s) + 2 \text{ H}_2\text{O}(l) \rightarrow 2 \text{ LiOH(aq)} + \text{H}_2(g)$$
$$\text{MgH}_2(s) + 2 \text{ H}_2\text{O}(l) \rightarrow \text{Mg(OH)}_2(s) + \text{H}_2(g)$$
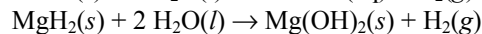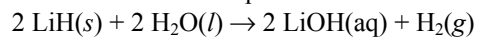
LiOH is water soluble, however, Mg(OH)_2 is not water soluble.

$V = nRT / P$

LiH $V =$

$$\frac{\left[1.00 \text{ lb LiH}\left(\frac{1 \text{ kg}}{2.205 \text{ lb}}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol LiH}}{7.946 \text{ g LiH}}\right)\left(\frac{1 \text{ mol H}_2}{2 \text{ mol LiH}}\right)\right]\left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 27)\text{K}\right)}{(750. \text{ torr})}\left(\frac{760 \text{ torr}}{1 \text{ atm}}\right)$$

= 712.213 = **712 L H$_2$ from LiH**

MgH$_2$ $V =$

$$\frac{\left[1.00 \text{ lb LiH}\left(\frac{1 \text{ kg}}{2.205 \text{ lb}}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol MgH}_2}{26.33 \text{ g MgH}_2}\right)\left(\frac{1 \text{ mol H}_2}{1 \text{ mol MgH}_2}\right)\right]\left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 27)\text{K}\right)}{(750.\text{torr})}\left(\frac{760 \text{ torr}}{1 \text{ atm}}\right)$$

= 429.870 = **4.30 x 10^2 L H$_2$ from MgH$_2$**

5.126   The root mean speed is related to temperature and molar mass as shown in equation:

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

$$u_{rms} \text{ Ne} = \sqrt{\frac{3(8.314\,\text{J/mol} \cdot \text{K})(370\text{ K})}{(20.18\text{ g/mol})}\left(\frac{10^3 \text{ g}}{\text{kg}}\right)\left(\frac{\text{kg} \cdot \text{m}^2/\text{s}^2}{\text{J}}\right)} = 676.24788 = \textbf{676 m/s Ne}$$

$$u_{rms} \text{ Ar} = \sqrt{\frac{3(8.314\,\text{J/mol} \cdot \text{K})(370\text{K})}{(39.95\text{ g/mol})}\left(\frac{10^3 \text{ g}}{\text{kg}}\right)\left(\frac{\text{kg} \cdot \text{m}^2/\text{s}^2}{\text{J}}\right)} = 480.6269 = \textbf{481 m/s Ar}$$

$$u_{rms} \text{ He} = \sqrt{\frac{3(8.314\,\text{J/mol} \cdot \text{K})(370\text{ K})}{(4.003\text{ g/mol})}\left(\frac{10^3 \text{ g}}{\text{kg}}\right)\left(\frac{\text{kg} \cdot \text{m}^2/\text{s}^2}{\text{J}}\right)} = 1518.356 = \textbf{1.52 x 10}^\textbf{3} \textbf{ m/s He}$$

5.127   a) Moles = $n = PV / RT$

$$\text{Moles} = \frac{(30.0 \text{ torr})(300 \text{ L})}{\left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273.2 + 37.0)\text{K}\right)}\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right) = 0.464990 \text{ mol (unrounded)}$$

Mass CO$_2$ = (0.464990 mol) (44.01 g/mol) = 20.4642 = **20.5 g CO$_2$**
Mass H$_2$O = (0.464990 mol) (18.02 g/mol) = 8.3791 = **8.38 g H$_2$O**
b) $\text{C}_6\text{H}_{12}\text{O}_6(s) + 6 \text{ O}_2(g) \rightarrow 6 \text{ CO}_2(g) + 6 \text{ H}_2\text{O}(g)$
Mass C$_6$H$_{12}$O$_6$ = (0.464990 mol CO$_2$/h) (1 mol C$_6$H$_{12}$O$_6$/6 mol CO$_2$) (180.16 g C$_6$H$_{12}$O$_6$ / 1 mol C$_6$H$_{12}$O$_6$) (8 h)
        = 111.6968 = 1 x 10^2 g C$_6$H$_{12}$O$_6$ (= body mass lost)
 (This assumes the significant figures are limited by the 8 h.)
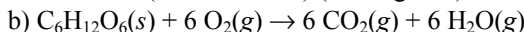
5.128 a) The number of moles of water gas can be found using the ideal gas equation, and moles times molar mass gives the mass of water.

$$\text{Moles } H_2O = \frac{(9.0 \text{ atm})(0.25 \text{ mL})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 175)\text{K})}\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{75\%}{100\%}\right)$$

$$= 4.58799 \times 10^{-5} \text{ mol } H_2O \text{ (unrounded)}$$

Mass =

$$\left(4.58799 \times 10^{-5} \text{ mol} H_2O\right)\left(\frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O}\right)\left(\frac{100\%}{1.5\%}\right) = 0.055117 = \mathbf{0.055 \text{ g}}$$

The last percent adjustment compensates for the water content being 1.5% of the mass.

b) $V = nRT / P =$

$$\frac{\left(4.58799 \times 10^{-5} \text{ mol } H_2O\right)\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 25)\text{K})}{1.00 \text{ atm}}\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$$

$$= 1.1224888 = \mathbf{1.1 \text{ mL}}$$

5.129 The balanced chemical equations are: $CaCO_3(s) + SO_2(g) \rightarrow CaSO_3(s) + CO_2(g)$
    $2 CaSO_3(s) + O_2(g) \rightarrow CaSO_4(s)$

Assuming the pressure and temperature from part b applies to part a:

Mole fraction $SO_2 = 1000 (2 \times 10^{-10}) = 2 \times 10^{-7}$

a) $\dfrac{(1.00 \text{ atm})(4 \text{ GL})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 25)\text{K})}\left(\dfrac{10^9 \text{ L}}{1 \text{ GL}}\right)\left(2 \times 10^{-7}\right)\left(\dfrac{95\%}{100\%}\right)\left(\dfrac{1 \text{ mol } CaSO_4}{1 \text{ mol } SO_2}\right)\left(\dfrac{136.15 \text{ g } CaSO_4}{1 \text{ mol } CaSO_4}\right)\left(\dfrac{1 \text{ kg}}{10^3 \text{ g}}\right)$

$$= 4.2293 = \mathbf{4 \text{ kg } CaSO_4}$$

b)

$$\left(4.2293 \text{ kg } CaSO_4\right)\left(\frac{1 \text{ mol } CaSO_4}{136.15 \text{ g } CaSO_4}\right)\left(\frac{1 \text{ mol } O_2}{2 \text{ mol } CaSO_4}\right)\left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273 + 25)\text{K})\left(\frac{1}{1.00 \text{ atm}}\right)\left(\frac{1}{0.209}\right)$$

$$= 1.8183 = \mathbf{2 \text{ L air}}$$

5.130 a) Derive $u_{rms} = \sqrt{\dfrac{3RT}{\mathcal{M}}}$

Set the given relationships equal to each other.

$$\frac{1}{2} m\overline{u^2} = \frac{3}{2}\left(\frac{R}{N_A}\right)T$$

$$\overline{u^2} = 3RT\left(\frac{R}{N_A}\right)T$$

$$\overline{u^2} = \frac{3RT}{mN_A} \text{ where } \mathcal{M} = mN_A$$

Solve for $\overline{u}^2$; substitute molar mass, $\mathcal{M}$, for mN_A (mass of one molecule x Avogadro's number of molecules).

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

b) Derive Graham's Law $\dfrac{\sqrt{\mathcal{M}_1}}{\sqrt{\mathcal{M}_2}} = \dfrac{rate_2}{rate_1}$

At a given T, the average kinetic energy for two substances, with molecular masses m_1 and m_2, is:

$$\overline{E}_k = \frac{1}{2} \, m_1\overline{u_1}^2 = \frac{1}{2} \, m_2\overline{u_2}^2$$

$$m_1\overline{u_1}^2 = m_2\overline{u_2}^2$$

$$\frac{m_1}{m_2} = \frac{\overline{u_2}^2}{\overline{u_1}^2} \qquad \longrightarrow \qquad \frac{\sqrt{m_1}}{\sqrt{m_2}} = \frac{\overline{u_2}}{\overline{u_1}}$$

Rearranging and taking the square root of both sides gives:

$$\frac{\sqrt{m_1}}{\sqrt{m_2}} = \frac{\overline{u_2}}{\overline{u_1}}$$

The average molecular speed, u, is directly proportional to the rate of effusion. Therefore, substitute "rate" for each "u." In addition, the molecular mass is directly proportional to the molar mass, so substitute $\mathcal{M}$ for each m:

$$\frac{\sqrt{\mathcal{M}_1}}{\sqrt{\mathcal{M}_2}} = \frac{rate_2}{rate_1}$$

5.131 The height of the column (measured parallel to the original vertical position of the tube) would remain constant, so the length of liquid column (measured along the no-longer-vertical tube) would increase. Part of the mass of the liquid column is now being supported by the glass of the tube, so its mass (and hence length) would be greater.

5.132 a) **cylinder B**   b) **cylinder B**   c) **none**   d) **cylinder C**   e) **cylinder D**

5.133 Ideal Gas Equation: PV = nRT

$$P_{0°} = nRT / V = \frac{\left[51.1 \text{ g NH}_3\left(\dfrac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3}\right)\right]\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 0)\text{K}\right)}{(3.000 \text{ L})}$$

$$= 22.417687 = \textbf{22.4 atm at 0°C}$$

$$P_{400°} = nRT / V = \frac{\left[51.1 \text{ g NH}_3\left(\dfrac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3}\right)\right]\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 400)\text{K}\right)}{(3.000 \text{ L})}$$

$$= 55.2641 = \textbf{55.3 atm at 400°C}$$

van der Waals Equation: $\left(P + \dfrac{n^2a}{V^2}\right)(V - nb) = nRT$

NH_3: a = 4.17 atm•L^2 / mol^2 b = 0.0371 L/mol

$$P_{VDW} = \frac{nRT}{V - nb} - \frac{n^2a}{V^2}$$

Moles NH_3 = (51.1 g NH_3) (1 mol NH_3 / 17.03 g NH_3) = 3.000587 mol NH_3 (unrounded)

$P_{0°}$ NH$_3$ =

$$\frac{(3.000587 \text{ mol NH}_3)\left(0.08206\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 0)\text{K}\right)}{3.000 \text{ L} - (3.000587 \text{ mol H}_2)\left(0.0371\frac{\text{L}}{\text{mol}}\right)} - \frac{(3.000587 \text{ mol NH}_3)^2\left(4.17\frac{\text{atm} \cdot \text{L}^2}{\text{mol}^2}\right)}{(3.000 \text{ L})^2}$$

= 19.0986 = **19.1 atm NH$_3$**

$P_{400°}$ NH$_3$ =

$$\frac{(3.000587 \text{ mol NH}_3)\left(0.08206\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 400)\text{K}\right)}{3.000 \text{ L} - (3.000587 \text{ mol H}_2)\left(0.0371\frac{\text{L}}{\text{mol}}\right)} - \frac{(3.000587 \text{ mol NH}_3)^2\left(4.17\frac{\text{atm} \cdot \text{L}^2}{\text{mol}^2}\right)}{(3.000 \text{ L})^2}$$

= 53.2243 = **53.2 atm NH$_3$**

5.134 $X_{methane} = 1.00 - X_{argon} - X_{helium} = 1.00 - 0.35 - 0.25 = 0.40$
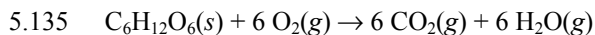$P_{methane} = X_{methane} P_{total} = (0.40) (1.75 \text{ atm}) = 0.70 \text{ atm CH}_4$
Moles x Avogadro's number = molecules
Moles = PV / RT

$$\text{Molecules} = \frac{(0.70 \text{ atm})(6.0 \text{ L})}{\left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 45)\text{K}\right)}\left(\frac{6.022 \times 10^{23} \text{ molecules CH}_4}{1 \text{ mol}}\right)$$

= 9.6876795 x 10^{22} = **9.7 x 10^{22} molecules CH$_4$**

5.135 $C_6H_{12}O_6(s) + 6 \text{ O}_2(g) \rightarrow 6 \text{ CO}_2(g) + 6 \text{ H}_2O(g)$
a) First, find the number of moles of carbon dioxide and use the T and P given to calculate volume from the ideal gas equation.
V =

$$\frac{\left[18.0 \text{ g Glucose}\left(\frac{1 \text{ mol Glucose}}{180.16 \text{ g Glucose}}\right)\left(\frac{6 \text{ mol CO}_2}{1 \text{ mol Glucose}}\right)\right]\left(0.0821\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)\left((273 + 35)\text{K}\right)}{(780. \text{ torr})}\left(\frac{760 \text{ torr}}{1 \text{ atm}}\right)$$

= 14.76992 = **14.8 Liters CO$_2$**

This solution assumes that partial pressure of O$_2$ does not interfere with the reaction conditions.
b) From the stoichiometric ratios, calculate the moles of each gas and then use Dalton's law of partial pressures to determine the pressure of each gas.
Moles $C_6H_{12}O_6$ = (9.0 g glucose) (1 mol glucose / 180.16 g glucose) = 0.049956 mol glucose (unrounded)
Moles of other substances = (0.049956 mol glucose) (6 mol / 1 mol glucose) = 0.299736 mol (unrounded)
At 35°C, the vapor pressure of water is 42.2 torr. No matter how much water is produced, the partial pressure of H$_2$O will still be 42.2 torr. The remaining pressure, 780 torr - 42.2 torr = 737.8 torr (unrounded) is the sum of partial pressures for O$_2$ and CO$_2$. Since the mole fractions of O$_2$ and CO$_2$ are equal, their pressures must be equal, and be one-half of sum of the partial pressures just found.
P$_{water}$ = 42.2 torr
(737.8 torr) / 2 = 368.9 = **3.7 x 10^2 torr P$_{oxygen}$ = P$_{carbon\ dioxide}$**

5.136   The average kinetic energies are the same for any gases at the same temperature:
Average kinetic energy = (3/2)RT = (3/2) (8.314 J/mol•K) (273 K)
= 3.40458 x 10^3 = **3.40 x 10^3 J** for both gases
rms speed:

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

$$u_{rms}N_2 = \sqrt{\frac{3(8.314\,\text{J/mol} \cdot \text{K})(273\,\text{K})}{(28.02\,\text{g/mol})}\left(\frac{10^3\,\text{g}}{\text{kg}}\right)\left(\frac{\text{kg} \cdot \text{m}^2/\text{s}^2}{\text{J}}\right)} = 4.92961\text{x }10^2 = \textbf{4.93 x 10}^2 \textbf{ m/s N}_2$$

$$u_{rms}\,H_2 = \sqrt{\frac{3(8.314\,\text{J/mol} \cdot \text{K})(273\,\text{K})}{(2.016\,\text{g/mol})}\left(\frac{10^3\,\text{g}}{\text{kg}}\right)\left(\frac{\text{kg} \cdot \text{m}^2/\text{s}^2}{\text{J}}\right)} = 1.83781\text{ x }10^3 = \textbf{1.84 x 10}^3 \textbf{ m/s H}_2$$

5.137 a) Assuming the total pressure is 760 torr.
       (0.2 torr Br_2 / 760 torr) (10^6) = 263.15789 = 300 ppm Br_2 **Unsafe**
   b) Assuming the total pressure is 760 torr.
       (0.2 torr CO_2 / 760 torr) (10^6) = 263.15789 = 300 ppm CO_2 **Safe**
   c) Moles bromine = (0.0004 g Br_2) (1 mol Br_2 / 159.80 g Br_2) = 2.5031 x 10^{-6} mol Br_2 (unrounded)

$$\text{Moles air} = PV / RT = \frac{(1.00\,\text{atm})(1000\,\text{L})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(273\,\text{K})} = 44.616\,\text{mol air (unrounded)}$$

       Concentration of Br_2 = [(2.5031 x 10^{-6} mol) / (44.616 mol)] (10^6) = 0.056103 = 0.06 ppm Br_2 **Safe**
   d) Moles CO_2 = (2.8 x 10^{22} molecules CO_2) (1 mol CO_2 / 6.022 x 10^{23} molecules CO_2)
       = 0.046496 mol CO_2 (unrounded)
Concentration of CO_2 = [(0.046496 mol) / (44.616 mol)] (10^6) = 1042.1 = 1.0 x 10^3 ppm CO_2 **Safe**

5.138 4 $NH_3(g)$ + 2 $NO(g)$ + 2 $O_2(g)$ → 3 $N_2(g)$ + 6 $H_2O(g)$

   a)
$$\frac{(4.5 \text{ x } 10^{-5}\,\text{atm})(1.00\,\text{L})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273+150)\text{K})}\left(\frac{4\text{ L }NH_3}{2\text{ L }NO}\right)$$

       = 2.5915 x 10^{-6} = **2.6 x 10**$^{-6}$ **L NH**$_3$ for every liter of polluted air

   b)
$$\frac{(4.5 \text{ x } 10^{-5}\,\text{atm})(\text{kL})}{\left(0.0821\dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)((273+150)\text{K})}\left(\frac{10^3\,\text{L}}{1\,\text{kL}}\right)\left(\frac{4\text{ mol }NH_3}{2\text{ mol }NO}\right)\left(\frac{17.03\text{ g }NH_3}{1\text{ mol }NH_3}\right)$$

       = 0.044134 = **0.044 g NH**$_3$

5.139 Rate Ne / Rate Xe = $\sqrt{\dfrac{\text{Molar Mass Xe}}{\text{Molar Mass Ne}}} = \sqrt{\dfrac{131.3\,\text{g/mol}}{20.18\,\text{g/mol}}}$ = 2.55077 enrichment factor (unrounded)

   Thus X_{Ne} = (2.55077) / (2.55077 + 1) = 0.7183709 = **0.7184**

5.140 To find the number of steps through the membrane, calculate the molar masses to find the ratio of effusion rates.
   This ratio is the enrichment factor for each step.
   From the margin note, $\mathcal{M}_{235}$ = 349.03 g/mol and $\mathcal{M}_{238}$ = 352.04 g/mol

   Rate ^{235}U / Rate ^{238}U = $\sqrt{\dfrac{\text{Molar Mass }^{238}UF_6}{\text{Molar Mass }^{235}UF_6}} = \sqrt{\dfrac{352.04\,\text{g/mol}}{349.03\,\text{g/mol}}}$

       = 1.004302694 enrichment factor (unrounded)
   Therefore, the abundance of $^{235}UF_6$ after one membrane is 0.72% x 1.004302694, and after "N" membranes:
   Abundance of $^{235}UF_6$ after "N" membranes = 0.72% * $(1.004302694)^N$
   Desired abundance of $^{235}UF_6$ = 3.0% = 0.72% * $(1.004302694)^N$
   Solving for N:
   4.16667 = $(1.004302694)^N$
   ln 4.16667 = ln $(1.004302694)^N$
   ln 4.16667 = N * ln (1.004302694)
   N = (ln 4.16667) / (ln 1.004302694) = 332.392957 = **332 steps**

5.141 $\left(P + \dfrac{n^2a}{V^2}\right)(V - nb) = nRT$

$P_{VDW} = \dfrac{nRT}{V - nb} - \dfrac{n^2a}{V^2}$

Use Table 5.5 to find the values of the van der Waals constants.
CO: $a = 1.45$ atm•L^2/mol^2 $\qquad$ $b = 0.0395$ L/mol

$$\dfrac{(1.000 \text{ mol CO})\left(0.08206\dfrac{L \bullet atm}{mol \bullet K}\right)(273.15 \text{ K})}{22.414 \text{ L} - (1.000 \text{ mol CO})\left(0.0395\dfrac{L}{mol}\right)} - \dfrac{(1.000 \text{ mol CO})^2\left(1.45\dfrac{atm \bullet L^2}{mol^2}\right)}{(22.414 \text{ L})^2}$$

$= 0.998909977 = 0.9989$ atm

5.142 The amount of each gas that leaks from the balloon is proportional to its effusion rate. Using 45% or 0.45 as the rate for H_2, the rate for O_2 can be determined from Graham's Law.

$$\text{Rate oxygen / Rate hydrogen} = \sqrt{\dfrac{\text{Molar Mass } H_2}{\text{Molar Mass } O_2}} = \sqrt{\dfrac{2.016 \text{ g/mol}}{32.00 \text{ g/mol}}} = \dfrac{\text{Rate } O_2}{\left(\dfrac{45\%}{100\%}\right)}$$

Rate $O_2 = 0.112949(100\%) = 11.2949$ (unrounded)
$O_2 / H_2 = (100 - 11.2949) / (100 - 45) = 1.6128 = \textbf{1.6}$

5.143 a) Options for PCl_3:
All values are g/mol

P	First Cl	Second Cl	Third Cl	Total
31	35	35	35	136
31	37	35	35	138
31	37	37	35	140
31	37	37	37	142

b) The fraction abundances are $^{35}Cl = 75\% / 100\% = 0.75$, and $^{37}Cl = 25\% / 100\% = 0.25$
The relative amount of each mass comes from the product of the relative abundances of each Cl isotope.

Mass 136 = (0.75) (0.75) (0.75) = 0.421875 = 0.42 (most abundant)
Mass 138 = (0.25) (0.75) (0.75) = 0.140625 = 0.14
Mass 140 = (0.25) (0.25) (0.75) = 0.046875 = 0.047
Mass 142 = (0.25) (0.25) (0.25) = 0.015625 = 0.016

c) Rate $P^{37}Cl_3$ / Rate $P^{35}Cl_3 = \sqrt{\dfrac{\text{Molar Mass } P^{35}Cl_3}{\text{Molar Mass } P^{37}Cl_3}} = \sqrt{\dfrac{136 \text{ g/mol}}{142 \text{ g/mol}}}$

$= 0.978645 = \textbf{0.979}$

5.144 a) $P_f = P_iT_f / T_i = (35.0 \text{ psi}) (319 \text{ K}) / (295 \text{ K}) = 37.847 = \textbf{37.8 psi}$
b) New tire volume = V_i (102% /100%) = 1.02 $V_i = V_f$
$\qquad P_f = P_iV_iT_f / T_iV_f = [(35.0 \text{ psi}) (V_i) (319 \text{ K})] / [(295 \text{ K}) (1.02 V_i)] = 37.105 = \textbf{37.1 psi}$

c) $\Delta n = \Delta PV / RT = \dfrac{((37.105 - 35.0)\text{psi})(208 \text{ L})}{\left(0.0821\dfrac{L \bullet atm}{mol \bullet K}\right)(295 \text{ K})}\left(\dfrac{1 \text{ atm}}{14.7 \text{ psi}}\right) = 1.229795579$ mol lost (unrounded)

Time = (1.229795579 mol) (28.8 g/mol) (1 min / 1.5 g) = 23.612 = **24 min**

5.145 a) d = $\mathcal{M}$ P / RT

$$O_2: d = \frac{(32.00 \text{ g/mol})(760 \text{ torr})}{\left(0.0821 \dfrac{L \bullet atm}{mol \bullet K}\right)((273 + 0)K)}\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right) = 1.42772 = \textbf{1.43 g O}_2 \textbf{ / L}$$

$$O_3: d = \frac{(48.00 \text{ g/mol})(760 \text{ torr})}{\left(0.0821 \dfrac{L \bullet atm}{mol \bullet K}\right)((273 + 0)K)}\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right) = 2.141585576 = \textbf{2.14 g O}_3 \textbf{ / L}$$

b) d_{ozone} / d_{oxygen} = (2.141585576) / (1.42772) = **1.5**
The ratio of the density is the same as the ratio of the molar masses.

5.146 The balanced chemical equation is: $7 F_2(g) + I_2(s) \rightarrow 2 IF_7(g)$
It will be necessary to determine the limiting reactant.

$$\text{Moles IF}_7 \text{ from F}_2 = \frac{(350. \text{ torr})(2.50 \text{ L})}{\left(0.0821\dfrac{L \bullet atm}{mol \bullet K}\right)(250. \text{ K})}\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right)\left(\frac{2 \text{ mol IF}_7}{7 \text{ mol F}_2}\right)$$

$= 0.016026668 \text{ mol IF}_7$ (unrounded)
Moles IF_7 from I_2 = (2.50 g I_2) (1 mol I_2 / 253.8 g I_2) (2 mol IF_7 / 1 mol I_2)
        $= 0.019700551 \text{ mol IF}_7$ (unrounded)
F_2 is limiting.

$$\text{Mole I}_2 \text{ remaining} = (2.50 \text{ g I}_2) (1 \text{ mol I}_2 / 253.8 \text{ g I}_2) - \frac{(350. \text{ torr})(2.50 \text{ L})}{\left(0.0821\dfrac{L \bullet atm}{mol \bullet K}\right)(250. \text{ K})}\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right)\left(\frac{1 \text{ mol I}_2}{7 \text{ mol F}_2}\right)$$

        $= 1.83694 \text{ x } 10^{-3} \text{ mol I}_2$ (unrounded)
Total moles of gas = (0 mol F_2) + (0.016026668 mol IF_7) + (1.83694 x 10^{-3} mol I_2)
        = 0.0178636 mol gas (unrounded)

$$P_{total} = nRT / V = \frac{(0.0178636 \text{ mol})\left(0.0821\dfrac{L \bullet atm}{mol \bullet K}\right)(550. \text{ K})}{(2.50 \text{ L})} = 0.322652 = \textbf{0.323 atm}$$

$P_{iodine} = X_{iodine} P_{total}$ = [(1.83694 x 10^{-3} mol I_2) / (0.0178636 mol)] (0.322652 atm) = 0.03317877 = **0.0332 atm**

5.147 a) n = PV / RT = $\dfrac{(85.0 \text{ atm})(850. \text{ L})}{\left(0.0821\dfrac{L \bullet atm}{mol \bullet K}\right)(298 \text{ K})}\left(\dfrac{80.0\%}{100\%}\right)$ = 2.36248 x 10^3 = **2.36 x 10^3 mol Cl$_2$**

b) $\left(P + \dfrac{n^2a}{V^2}\right)(V - nb) = nRT$

$$P_{VDW} = \frac{nRT}{V - nb} - \frac{n^2a}{V^2}$$

Use Table 5.5 to find the values of the van der Waals constants.
Cl_2: a = 6.49 atm•L^2/mol^2 b = 0.0562 L/mol

$$\frac{(2.36248 \text{ x } 10^3 \text{ mol Cl}_2)\left(0.08206\dfrac{L \bullet atm}{mol \bullet K}\right)(298 \text{ K})}{850. \text{ L} - (2.36248 \text{ x } 10^3 \text{ mol Cl}_2)\left(0.0562\dfrac{L}{mol}\right)} - \frac{(2.36248 \text{ x } 10^3 \text{ mol Cl}_2)^2\left(6.49\dfrac{atm \bullet L^2}{mol^2}\right)}{(850. \text{ L})^2}$$

        = 30.4134 = **30.4 atm**
c) The engineer did not completely fill the tank. She should have filled it to (80.0%/100%) (85.0 atm) = 68 atm, but only filled it to 30.4 atm.