

CHAPTER 6 GASES

PRACTICE EXAMPLES

1A The pressure measured by each liquid must be the same. They are related through $P = g h d$. Thus, we have the following $g h_{\text{DEG}} d_{\text{DEG}} = g h_{\text{Hg}} d_{\text{Hg}}$. The g 's cancel; we substitute known values: $9.25 \text{ m}_{\text{DEG}} \times 1.118 \text{ g/cm}^3_{\text{DEG}} = h_{\text{Hg}} \times 13.6 \text{ g/cm}^3_{\text{Hg}}$

$$h_{\text{Hg}} = 9.25 \text{ m} \times \frac{1.118 \text{ g/cm}^3}{13.6 \text{ g/cm}^3} = 0.760 \text{ m Hg}, \quad P = 0.760 \text{ m Hg} = 760. \text{ mmHg}$$

1B The solution is found through the expression relating density and height: $h_{\text{TEG}} d_{\text{TEG}} = h_{\text{Hg}} d_{\text{Hg}}$

We substitute known values and solve for triethylene glycol's density:

$9.14 \text{ m}_{\text{TEG}} \times d_{\text{TEG}} = 757 \text{ mmHg} \times 13.6 \text{ g/cm}^3_{\text{Hg}}$. Using unit conversions, we get

$$d_{\text{TEG}} = \frac{0.757 \text{ m}}{9.14 \text{ m}} \times 13.6 \text{ g/cm}^3 = 1.13 \text{ g/cm}^3$$

2A We know that $P_{\text{gas}} = P_{\text{bar}} + \Delta P$ with $P_{\text{bar}} = 748.2 \text{ mmHg}$. We are told that $\Delta P = 7.8 \text{ mmHg}$. Thus, $P_{\text{gas}} = 748.2 \text{ mmHg} + 7.8 \text{ mmHg} = 756.0 \text{ mmHg}$.

2B The difference in pressure between the two levels must be the same, just expressed in different units. Hence, this problem is almost a repetition of Practice Example 6-1.

$h_{\text{Hg}} = 748.2 \text{ mmHg} - 739.6 \text{ mmHg} = 8.6 \text{ mmHg}$. Again we have $g h_g d_g = g h_{\text{Hg}} d_{\text{Hg}}$.

This becomes $h_g \times 1.26 \text{ g/cm}^3_{\text{glycerol}} = 8.6 \text{ mmHg} \times 13.6 \text{ g/cm}^3_{\text{Hg}}$

$$h_g = 8.6 \text{ mmHg} \times \frac{13.6 \text{ g/cm}^3_{\text{Hg}}}{1.26 \text{ g/cm}^3_{\text{glycerol}}} = 93 \text{ mm glycerol}$$

3A $A = \pi r^2$ (here $r = \frac{1}{2}(2.60 \text{ cm} \times \frac{1 \text{ m}}{100 \text{ cm}}) = 0.0130 \text{ m}$)

$$A = \pi(0.0130 \text{ m})^2 = 5.31 \times 10^{-4} \text{ m}^2$$

$$F = m \times g = (1.000 \text{ kg})(9.81 \text{ m s}^{-2}) = 9.81 \text{ kg m s}^{-2} = 9.81 \text{ N}$$

$$P = \frac{F}{A} = \frac{9.81 \text{ N}}{5.31 \times 10^{-4} \text{ m}^2} = 18475 \text{ N m}^{-2} \text{ or } 1.85 \times 10^4 \text{ Pa}$$

$$P (\text{torr}) = 1.85 \times 10^4 \text{ Pa} \times = 139 \text{ torr}$$

3B Final pressure = 100 mb. $100 \text{ mb} \times \frac{101,325 \text{ Pa}}{1013.25 \text{ mb}} = 1.000 \times 10^4 \text{ Pa}$

The area of the cylinder is unchanged from that in Example 6-3, $(1.32 \times 10^{-3} \text{ m}^2)$.

$$P = \frac{F}{A} = 1.000 \times 10^4 \text{ Pa} = \frac{F}{1.32 \times 10^{-3} \text{ m}^2}$$

Solving for F, we find $F = 13.2 \text{ (Pa)m}^2 = 13.2 \text{ (N m}^{-2}\text{)m}^2 = 13.2 \text{ N}$

$$F = m \times g = 13.2 \text{ kg m s}^{-2} = m \times 9.81 \text{ m s}^{-2}$$

$$\text{total mass} = \text{mass of cylinder} + \text{mass added weight} = m = \frac{F}{g} = \frac{13.2 \text{ kg m s}^{-2}}{9.81 \text{ m s}^{-2}} = 1.35 \text{ kg}$$

An additional 350 grams must be added to the top of the 1.000 kg (1000 g) red cylinder to increase the pressure to 100 mb. It is not necessary to add a mass with the same cross sectional area. The pressure will only be exerted over the area that is the base of the cylinder on the surface beneath it.

4A The ideal gas equation is solved for volume. Conversions are made within the equation.

$$V = \frac{nRT}{P} = \frac{\left(20.2 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3}\right) \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times (-25 + 273) \text{ K}}{752 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}} = 24.4 \text{ L NH}_3$$

4B The amount of $\text{Cl}_2(\text{g})$ is 0.193 mol Cl_2 and the pressure is 0.980 atm (0.993 barr \times (1 atm/1.01325 barr) = 0.980 atm). This information is substituted into the ideal gas equation after it has been solved for temperature.

$$T = \frac{PV}{nR} = \frac{0.980 \text{ atm} \times 7.50 \text{ L}}{0.193 \text{ mol} \times 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1}} = 464 \text{ K}$$

5A The ideal gas equation is solved for amount and the quantities are substituted.

$$n = \frac{PV}{RT} = \frac{10.5 \text{ atm} \times 5.00 \text{ L}}{\frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times (30.0 + 273.15) \text{ K}} = 2.11 \text{ mol He}$$

5B

$$n = \frac{PV}{RT} = \frac{\left(6.67 \times 10^{-7} \text{ Pa} \times \frac{1 \text{ atm}}{101325 \text{ Pa}}\right) \left(3.45 \text{ m}^3 \times \frac{1000 \text{ L}}{1 \text{ m}^3}\right)}{(0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}) (25 + 273.15) \text{ K}} = 9.28 \times 10^{-10} \text{ moles of N}_2$$

$$\text{molecules of N}_2 = 9.28 \times 10^{-10} \text{ mol N}_2 \times \frac{6.022 \times 10^{23} \text{ molecules of N}_2}{1 \text{ mole N}_2}$$

$$\text{molecules of N}_2 = 5.59 \times 10^{14} \text{ molecules N}_2$$

6A The general gas equation is solved for volume, after the constant amount in moles is cancelled. Temperatures are converted to kelvin.

$$V_2 = \frac{V_1 P_1 T_2}{P_2 T_1} = \frac{1.00 \text{ mL} \times 2.14 \text{ atm} \times (37.8 + 273.2) \text{ K}}{1.02 \text{ atm} \times (36.2 + 273.2) \text{ K}} = 2.11 \text{ mL}$$

6B The flask has a volume of 1.00 L and initially contains $O_2(g)$ at STP. The mass of $O_2(g)$ that must be released is obtained from the difference in the amount of $O_2(g)$ at the two temperatures, 273 K and 373 K. We also could compute the masses separately and subtract them. We note that 1.00 bar is 0.987 atm.

$$\begin{aligned} \text{mass released} &= (n_{\text{STP}} - n_{100^\circ\text{C}}) \times M_{O_2} = \left(\frac{PV}{R273\text{ K}} - \frac{PV}{R373\text{ K}} \right) \times M_{O_2} = \frac{PV}{R} \left(\frac{1}{273\text{ K}} - \frac{1}{373\text{ K}} \right) \times M_{O_2} \\ &= \frac{0.987\text{ atm} \times 1.00\text{ L}}{0.08206\text{ L atm mol}^{-1}\text{ K}^{-1}} \left(\frac{1}{273\text{ K}} - \frac{1}{373\text{ K}} \right) \times \frac{32.00\text{ g}}{1\text{ mol } O_2} = 0.378\text{ g } O_2 \end{aligned}$$

7A The volume of the vessel is 0.09841 L. We substitute other values into the expression for molar mass.

$$M = \frac{mRT}{PV} = \frac{(40.4868\text{ g} - 40.1305\text{ g}) \times \frac{0.08206\text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times (22.4 + 273.2)\text{ K}}{\left(772\text{ mmHg} \times \frac{1\text{ atm}}{760\text{ mmHg}} \right) \times 0.09841\text{ L}} = 86.4\text{ g/mol}$$

7B The gas's molar mass is its mass (1.27 g) divided by the amount of the gas in moles. The amount can be determined from the ideal gas equation.

$$\begin{aligned} n &= \frac{PV}{RT} = \frac{\left(737\text{ mmHg} \times \frac{1\text{ atm}}{760\text{ mmHg}} \right) \times 1.07\text{ L}}{\frac{0.08206\text{ L atm}}{\text{mol} \cdot \text{K}} \times (25 + 273)\text{ K}} = 0.0424\text{ mol gas} \\ M &= \frac{1.27\text{ g}}{0.0424\text{ mol}} = 30.0\text{ g/mol} \end{aligned}$$

This answer is in good agreement with the molar mass of NO, 30.006 g/mol.

8A The molar mass of He is 4.003 g/mol. This is substituted into the expression for density.

$$d = \frac{MP}{RT} = \frac{4.003\text{ g mol}^{-1} \times 0.987\text{ atm}}{0.08206\text{ L} \cdot \text{atm mol}^{-1}\text{ K}^{-1} \times 298\text{ K}} = 0.162\text{ g/L}$$

When compared to the density of air under the same conditions (1.16 g/L, based on the "average molar mass of air" = 28.8 g/mol) the density of He is only about one seventh as much. Thus, helium is less dense ("lighter") than air.

8B The suggested solution is a simple one; we merely need to solve for mass of gas from density and its moles from the ideal gas law.

$$\begin{aligned} m(\text{gas}) &= D \times V = (1.00\text{ g/L})(1.00\text{ L}) = 1.00\text{ g} \\ n &= \frac{PV}{RT} = \frac{\left(745\text{ mmHg} \times \frac{1\text{ atm}}{760\text{ mmHg}} \right) (1.00\text{ L})}{0.08206\text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \times 382\text{ K}} = 0.0312\text{ mol} \end{aligned}$$

Therefore, the molar mass of the gas is as follows:

$$M = \frac{\text{g}}{\text{mol}} = \frac{1.00 \text{ g}}{0.0312 \text{ mol}} = 32.0 \text{ g/mol}$$

The molecular weight suggests that the gas is O_2 .

9A The balanced equation is $2 \text{NaN}_3(\text{s}) \xrightarrow{\Delta} 2 \text{Na}(\text{l}) + 3 \text{N}_2(\text{g})$

$$\text{moles N}_2 = \frac{P V}{R T} = \frac{\left(776 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}\right) \times 20.0 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{mol K}} \times (30.0 + 273.2) \text{ K}} = 0.821 \text{ mol N}_2$$

Now, solve the stoichiometry problem.

$$\text{mass NaN}_3 = 0.821 \text{ mol N}_2 \times \frac{2 \text{ mol NaN}_3}{3 \text{ mol N}_2} \times \frac{65.01 \text{ g NaN}_3}{1 \text{ mol NaN}_3} = 35.6 \text{ g NaN}_3$$

9B Here we are not dealing with gaseous reactants; the law of combining volumes cannot be used. From the ideal gas equation we determine the amount of $\text{N}_2(\text{g})$ per liter under the specified conditions. Then we determine the amount of $\text{Na}(\text{l})$ produced simultaneously, and finally the mass of that $\text{Na}(\text{l})$.

$$\text{Pressure: } 1.0 \text{ barr} \times \frac{1 \text{ atm}}{1.01325 \text{ barr}} = 0.987 \text{ atm}$$

$$\text{mass of Na}(\text{l}) = \frac{0.987 \text{ atm} \times 1.000 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{mol K}} \times (25 + 273) \text{ K}} \times \frac{2 \text{ mol Na}}{3 \text{ mol N}_2} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 0.619 \text{ g Na}(\text{l})$$

10A The law of combining volumes permits us to use stoichiometric coefficients for volume ratios.

$$\text{O}_2 \text{ volume} = 1.00 \text{ L NO}(\text{g}) \times \frac{5 \text{ L O}_2}{4 \text{ L NO}} = 1.25 \text{ L O}_2(\text{g})$$

10B The first task is to balance the chemical equation. There must be three moles of hydrogen for every mole of nitrogen in both products (because of the formula of NH_3) and reactants: $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightarrow 2 \text{NH}_3(\text{g})$. The volumes of gaseous reactants and products are related by their stoichiometric coefficients, as long as all gases are at the same temperature and pressure.

$$\text{volume NH}_3(\text{g}) = 225 \text{ L H}_2(\text{g}) \times \frac{2 \text{ L NH}_3(\text{g})}{3 \text{ L H}_2(\text{g})} = 150. \text{ L NH}_3$$

11A We can work easily with the ideal gas equation, with a the new temperature of $T = (55 + 273) \text{ K} = 328 \text{ K}$. The amount of Ne added is readily computed.

$$n_{\text{Ne}} = 12.5 \text{ g Ne} \times \frac{1 \text{ mol Ne}}{20.18 \text{ g Ne}} = 0.619 \text{ mol Ne}$$

$$P = \frac{n_{\text{total}}RT}{V} = \frac{(1.75 + 0.619)\text{mol} \times \frac{0.08206 \text{ L atm}}{\text{mol K}} \times 328 \text{ K}}{5.0 \text{ L}} = 13 \text{ atm}$$

11B The total volume initially is 2.0 L + 8.0 L = 10.0 L. These two mixed ideal gases then obey the general gas equation as if they were one gas.

$$P_2 = \frac{P_1V_1T_2}{V_2T_1} = \frac{1.00 \text{ atm} \times 10.0 \text{ L} \times 298 \text{ K}}{2.0 \text{ L} \times 273 \text{ K}} = 5.5 \text{ atm}$$

12A The partial pressures are proportional to the mole fractions.

$$P_{\text{H}_2\text{O}} = \frac{n_{\text{H}_2\text{O}}}{n_{\text{tot}}} \times P_{\text{tot}} = \frac{0.00278 \text{ mol H}_2\text{O}}{0.197 \text{ mol CO}_2 + 0.00278 \text{ mol H}_2\text{O}} \times 2.50 \text{ atm} = 0.0348 \text{ atm H}_2\text{O(g)}$$

$$P_{\text{CO}_2} = P_{\text{tot}} - P_{\text{H}_2\text{O}} = 2.50 \text{ atm} - 0.0348 \text{ atm} = 2.47 \text{ atm CO}_2(\text{g})$$

12B Expression (6.17) indicates that, in a mixture of gases, the mole percent equals the volume percent, which in turn equals the pressure percent. Thus, we can apply these volume percents—converted to fractions by dividing by 100—directly to the total pressure.

$$\text{N}_2 \text{ pressure} = 0.7808 \times 748 \text{ mmHg} = 584 \text{ mmHg,}$$

$$\text{O}_2 \text{ pressure} = 0.2095 \times 748 \text{ mmHg} = 157 \text{ mmHg,}$$

$$\text{CO}_2 \text{ pressure} = 0.00036 \times 748 \text{ mmHg} = 0.27 \text{ mmHg,}$$

$$\text{Ar pressure} = 0.0093 \times 748 \text{ mmHg} = 7.0 \text{ mmHg}$$

13A First compute the moles of $\text{H}_2(\text{g})$, then use stoichiometry to convert to moles of HCl.

$$\text{amount HCl} = \frac{\left((755 - 25.2) \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} \right) \times 0.0355 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{mol K}} \times (26 + 273) \text{ K}} \times \frac{6 \text{ mol HCl}}{3 \text{ mol H}_2} = 0.00278 \text{ mol HCl}$$

13B The volume occupied by the $\text{O}_2(\text{g})$ at its partial pressure is the same as the volume occupied by the mixed gases: water vapor and $\text{O}_2(\text{g})$. The partial pressure of $\text{O}_2(\text{g})$ is found by difference.

$$\text{O}_2 \text{ pressure} = 749.2 \text{ total pressure} - 23.8 \text{ mmHg (H}_2\text{O pressure)} = 725.4 \text{ mmHg}$$

$$\text{mol O}_2 = \frac{P \cdot V}{R \cdot T} = \frac{\left(725.4 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} \right) (0.395 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1})(298 \text{ K})} = 0.0154 \text{ mol O}_2$$

$$\text{mass Ag}_2\text{O} = 0.0154 \text{ mol O}_2 \times \frac{2 \text{ mol Ag}_2\text{O}}{1 \text{ mol O}_2} \times \frac{231.74 \text{ g Ag}_2\text{O}}{1 \text{ mol Ag}_2\text{O}} = 7.14 \text{ g Ag}_2\text{O}$$

$$\text{Mass\% Ag}_2\text{O} = 7.14/8.07 \times 100 = 88.4\%$$

The volume of the dry gas in turn is determined as follows:

$$V = \frac{(0.0154 \text{ mol})(0.08206 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1})(298 \text{ K})}{(749.2 / 760 \text{ atm})} = 0.382 \text{ L}$$

14A The gas with the smaller molar mass, NH_3 at 17.0 g/mol, has the greater root-mean-square speed

$$u_{\text{rms}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3 \times 8.3145 \text{ kg m}^2 \text{ s}^{-2} \text{ mol}^{-1} \text{ K}^{-1} \times 298 \text{ K}}{0.0170 \text{ kg mol}^{-1}}} = 661 \text{ m/s}$$

14B bullet speed = $\frac{2180 \text{ mi}}{1 \text{ h}} \times \frac{1 \text{ h}}{3600 \text{ s}} \times \frac{5280 \text{ ft}}{1 \text{ mi}} \times \frac{12 \text{ in.}}{1 \text{ ft}} \times \frac{2.54 \text{ cm}}{1 \text{ in.}} \times \frac{1 \text{ m}}{100 \text{ cm}} = 974.5 \text{ m/s}$

Solve the rms-speed equation (6.20) for temperature by first squaring both sides.

$$(u_{\text{rms}})^2 = \frac{3RT}{M} \quad T = \frac{(u_{\text{rms}})^2 M}{3R} = \frac{\left(\frac{974.5 \text{ m}}{1 \text{ s}}\right)^2 \times \frac{2.016 \times 10^{-3} \text{ kg}}{1 \text{ mol H}_2}}{3 \times \frac{8.3145 \text{ kg m}^2}{\text{s}^2 \text{ mol K}}} = 76.75 \text{ K}$$

We expected the temperature to be lower than 298 K. Note that the speed of the bullet is about half the speed of a H_2 molecule at 298 K. To halve the speed of a molecule, its temperature must be divided by four.

15A The only difference is the gas's molar mass. 2.2×10^{-4} mol N_2 effuses through the orifice in 105 s.

$$\frac{? \text{ mol O}_2}{2.2 \times 10^{-4} \text{ mol N}_2} = \sqrt{\frac{M_{\text{N}_2}}{M_{\text{O}_2}}} = \sqrt{\frac{28.014 \text{ g/mol}}{31.999 \text{ g/mol}}} = 0.9357$$

$$\text{moles O}_2 = 0.9357 \times (2.2 \times 10^{-4}) = 2.1 \times 10^{-4} \text{ mol O}_2$$

15B Rates of effusion are related by the square root of the ratio of the molar masses of the two gases. H_2 , effuses faster (by virtue of being lighter), and thus requires a shorter time for the same amount of gas to effuse.

$$\text{time}_{\text{H}_2} = \text{time}_{\text{N}_2} \times \sqrt{\frac{M_{\text{H}_2}}{M_{\text{N}_2}}} = 105 \text{ s} \times \sqrt{\frac{2.016 \text{ g H}_2/\text{mol H}_2}{28.014 \text{ g N}_2/\text{mol N}_2}} = 28.2 \text{ s}$$

16A Effusion times are related as the square root of the molar mass. It requires 87.3 s for Kr to effuse.

$$\frac{\text{unknown time}}{\text{Kr time}} = \sqrt{\frac{M_{\text{unk}}}{M_{\text{Kr}}}} \quad \text{substitute in values} \quad \frac{131.3 \text{ s}}{87.3 \text{ s}} = \sqrt{\frac{M_{\text{unk}}}{83.80 \text{ g/mol}}} = 1.50$$

$$M_{\text{unk}} = (1.504)^2 \times 83.80 \text{ g/mol} = 1.90 \times 10^2 \text{ g/mol}$$

16B This problem is solved in virtually the same manner as Practice Example 18B. The lighter gas is ethane, with a molar mass of 30.07 g/mol.

$$\text{C}_2\text{H}_6 \text{ time} = \text{Kr time} \times \sqrt{\frac{M(\text{C}_2\text{H}_6)}{M(\text{Kr})}} = 87.3 \text{ s} \times \sqrt{\frac{30.07 \text{ g C}_2\text{H}_6 / \text{mol C}_2\text{H}_6}{83.80 \text{ g Kr} / \text{mol Kr}}} = 52.3 \text{ s}$$

17A Because one mole of gas is being considered, the value of n^2a is numerically the same as the value of a , and the value of nb is numerically the same as the value of b .

$$P = \frac{nRT}{V - nb} - \frac{n^2a}{V^2} = \frac{1.00 \text{ mol} \times \frac{0.083145 \text{ L barr}}{\text{mol K}} \times 273 \text{ K}}{(2.00 - 0.0427) \text{ L}} - \frac{3.66 \text{ L}^2 \text{ barr mol}^{-2}}{(2.00 \text{ L})^2} = 11.59 \text{ barr} - 0.915 \text{ barr}$$

$$= 10.68 \text{ barr CO}_2(\text{g}) \quad \text{compared with } 10.03 \text{ barr for Cl}_2(\text{g})$$

$$P_{\text{ideal}} = \frac{nRT}{V} = \frac{1.00 \text{ mol} \times \frac{0.083145 \text{ L barr}}{\text{mol K}} \times 273 \text{ K}}{2.00 \text{ L}} = 11.35 \sim 11.4 \text{ barr}$$

$\text{Cl}_2(\text{g})$ shows a greater deviation from ideal gas behavior than does $\text{CO}_2(\text{g})$.

17B Because one mole of gas is being considered, the value of n^2a is numerically the same as the value of a , and the value of nb is numerically the same as the value of b .

$$P = \frac{nRT}{V - nb} - \frac{n^2a}{V^2} = \frac{1.00 \text{ mol} \times \frac{0.083145 \text{ L barr}}{\text{mol K}} \times 273 \text{ K}}{(2.00 - 0.0395) \text{ L}} - \frac{1.47 \text{ L}^2 \text{ barr}}{(2.00 \text{ L})^2} = 11.58 \text{ atm} - 0.368 \text{ atm}$$

$$= 11.2 \text{ barr CO}(\text{g})$$

compared to 10.03 barr for $\text{Cl}_2(\text{g})$, 11.2 barr for CO, and 11.35 barr for $\text{CO}_2(\text{g})$.

Thus, $\text{Cl}_2(\text{g})$ displays the greatest deviation from ideality, 11.4 barr.

INTEGRATIVE EXERCISE

A. First, convert the available data to easier units.

$$101.3 \times 10^3 \text{ Pa} \times (1 \text{ atm} / 1.013 \times 10^5 \text{ Pa}) = 1 \text{ atm}, \text{ and } 25 \text{ }^\circ\text{C} = 298 \text{ K}.$$

Then, assume that we have a 1 L container of this gas, and determine how many moles of gas are present:

$$n = \frac{PV}{RT} = \frac{(1.00 \text{ atm})(1.00 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1})(298 \text{ K})} = 0.0409 \text{ mol}.$$

Knowing the density of the gas (1.637 g/L) and its volume (1 L) gives us the mass of 1 L of gas, or 1.637 g. Therefore, the molar mass of this gas is:

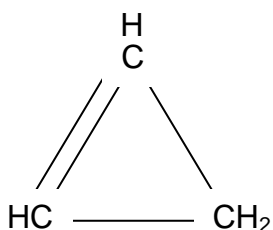
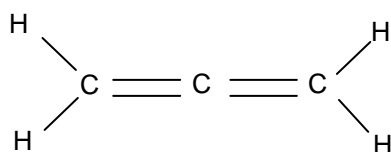
$$(1.637 \text{ g} / 0.0409 \text{ mol}) = 40.03 \text{ g/mol}$$

Now, determine the number of moles of C and H to ascertain the empirical formula:

$$\text{mol C: } 1.687 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.0383 \text{ mol}$$

$$\text{mol H: } 0.4605 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.05111 \text{ mol}$$

Dividing by the smallest value (0.0383 mol C), we get a H:C ratio of 1.33:1, or 4:3. Therefore, the empirical formula is C_3H_4 , which has a molar mass of 40.07, which is essentially the same as the molar mass calculated. Therefore, the actual formula is also C_3H_4 . Below are three possible Lewis structures:



B. First, let us determine the amount of each element in the compound:

$$\text{mol C: } 151.2 \times 10^{-3} \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.003435 \text{ mol C}$$

$$\text{g C: } 0.003435 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 0.04126 \text{ g C}$$

$$\text{mol H: } 69.62 \times 10^{-3} \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.007727 \text{ mol H}$$

$$\text{g H: } 0.007727 \text{ mol H} \times \frac{1.01 \text{ g H}}{1 \text{ mol H}} = 0.007804 \text{ g H}$$

$$\text{mol N: } \frac{PV}{RT} = \frac{1 \text{ atm} \cdot 9.62 \times 10^{-3} \text{ L}}{(0.08206 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1})(273 \text{ K})} \times \frac{2 \text{ mol N}}{1 \text{ mol N}_2} = 0.0008589 \text{ mol N}$$

$$\text{g N: } 0.0008589 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 0.01203 \text{ g N}$$

Therefore, the mass of O is determined by subtracting the sum of the above masses from the mass of the compound:

$$\text{g O: } 0.1023 - (0.04126 + 0.007804 + 0.01203) = 0.04121 \text{ g O}$$

$$\text{mol O: } 0.04121 \text{ g O} \times \frac{1 \text{ mol O}}{16.0 \text{ g O}} = 0.002576 \text{ mol O}$$

To determine the empirical formula, all of the calculated moles above should be divided by the smallest value. Doing so will give the following ratios:

$$\text{C: } 0.003435/0.0008589 = 4$$

$$\text{H: } 0.007727/0.0008589 = 9$$

$$\text{N: } 1$$

$$\text{O: } 0.002576/0.0008589 = 3$$

The empirical formula is $\text{C}_4\text{H}_9\text{NO}_3$, and has a molar mass of 119.14 g/mol.

To determine the actual formula, we have to calculate its molecular mass. We know the density at a given volume and therefore we need to find out the number of moles. Before that, we should convert the experimental conditions to more convenient units. $T = 127^\circ\text{C} + 273 = 400 \text{ K}$, and $P = 748 \text{ mm Hg}/760 \text{ mm Hg} = 0.9842 \text{ atm}$.

$$n = \frac{PV}{RT} = \frac{(0.9842 \text{ atm}) \cdot 1 \text{ L}}{(0.08205 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1})(400 \text{ K})} = 0.02999 \text{ mol}$$

$$\text{MM} = \text{g/mol} = 3.57 \text{ g}/0.02999 = 119 \text{ g/mol}$$

Therefore, the empirical and molecular formulas are the same, being $\text{C}_4\text{H}_9\text{NO}_3$.

EXERCISES

Pressure and Its Measurement

1. (a) $P = 736 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.968 \text{ atm}$

(b) $P = 0.776 \text{ bar} \times \frac{1 \text{ atm}}{1.01325 \text{ bar}} = 0.766 \text{ atm}$

(c) $P = 892 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.17 \text{ atm}$

(d) $P = 225 \text{ kPa} \times \frac{1000 \text{ Pa}}{1 \text{ kPa}} \times \frac{1 \text{ atm}}{101,325 \text{ Pa}} = 2.22 \text{ atm}$

3. We use: $h_{\text{bnz}} d_{\text{bnz}} = h_{\text{Hg}} d_{\text{Hg}}$

$$h_{\text{bnz}} = 0.970 \text{ atm} \times \frac{0.760 \text{ m Hg}}{1 \text{ atm}} \times \frac{13.6 \text{ g/cm}^3 \text{ Hg}}{0.879 \text{ g/cm}^3 \text{ benzene}} = 11.4 \text{ m benzene}$$

$$\underline{5.} \quad P = P_{\text{bar}} - h_1 = 740 \text{ mm Hg} - 30 \text{ mm (} h_1 \text{)} = 710 \text{ mm Hg}$$

$$\underline{7.} \quad F = m \times g \quad \text{and} \quad 1 \text{ atm} = 101325 \text{ Pa} = 101325 \text{ kg m}^{-1} \text{ s}^{-2} = P = \frac{F}{A} = \frac{m \times 9.81 \text{ m s}^{-2}}{1 \text{ m}^2}$$

$$\text{mass (per m}^2\text{)} = \frac{101325 \text{ kg m}^{-1} \text{ s}^{-2} \times 1 \text{ m}^2}{9.81 \text{ m s}^{-2}} = 10329 \text{ kg}$$

$$\text{(Note: } 1 \text{ m}^2 = (100 \text{ cm})^2 = 10,000 \text{ cm}^2\text{)}$$

$$P \text{ (kg cm}^{-2}\text{)} = \frac{m}{A} = \frac{10329 \text{ kg}}{10,000 \text{ cm}^2} = 1.03 \text{ kg cm}^{-2}$$

The Simple Gas Laws

$$\underline{9.} \quad \text{(a)} \quad V = 26.7 \text{ L} \times \frac{762 \text{ mmHg}}{385 \text{ mmHg}} = 52.8 \text{ L}$$

$$\text{(b)} \quad V = 26.7 \text{ L} \times \frac{762 \text{ mmHg}}{3.68 \text{ atm} \times \frac{760 \text{ mmHg}}{1 \text{ atm}}} = 7.27 \text{ L}$$

11. Charles' Law states that $V_1/T_1 = V_2/T_2$. Therefore,

$$\frac{3.0 \text{ L}}{450 \text{ K}} = \frac{1.50 \text{ L}}{T_2}, \quad \text{and} \quad T_2 = 225 \text{ K}$$

$$\underline{13.} \quad P_i = P_f \times \frac{V_f}{V_i} = \left(721 \text{ mmHg} \times \frac{35.8 \text{ L} + 1875 \text{ L}}{35.8 \text{ L}} \right) \times \frac{1 \text{ atm}}{760 \text{ mm H}_2\text{O}} = 50.6 \text{ atm}$$

15. Combining Boyle's and Charles' Law, we get the following expression:

$$\frac{P_1 \cdot V_1}{T_1} = \frac{P_2 \cdot V_2}{T_2}$$

Therefore,

$$T_2 = \frac{P_2 \cdot V_2 \cdot T_1}{P_1 \cdot V_1} = \frac{(0.340 \text{ atm})(5.00 \times 10^3 \text{ m}^3)(300 \text{ K})}{(1.000 \text{ atm})(2.00 \times 10^3 \text{ m}^3)} = 255 \text{ K}$$

17. STP: $P = 1 \text{ barr}$ and $T = 273.15$. $P (1 \text{ barr}) = 0.9869 \text{ atm}$

$$\text{mass}_{\text{Ar}} = \frac{0.9869 \text{ atm} \times 0.0750 \text{ L}}{0.08206 \frac{\text{L atm}}{\text{K mol}} \times 273.15 \text{ K}} \times \frac{39.948 \text{ g Ar}}{1 \text{ mol Ar}} = 0.132 \text{ g Ar}$$

- 19.** (a) *Conversion pathway approach:*

$$\begin{aligned} \text{mass} &= 27.6 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1 \text{ mol PH}_3}{22.698 \text{ L STP}} \times \frac{34.0 \text{ g PH}_3}{1 \text{ mol PH}_3} \times \frac{1000 \text{ mg PH}_3}{1 \text{ g}} \\ &= 41.3 \text{ mg PH}_3 \end{aligned}$$

Stepwise approach:

$$27.6 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.0276 \text{ L}$$

$$0.0276 \text{ L} \times \frac{1 \text{ mol PH}_3}{22.698 \text{ L STP}} = 0.001216 \text{ mol PH}_3$$

$$0.001216 \text{ mol PH}_3 \times \frac{34.0 \text{ g PH}_3}{1 \text{ mol PH}_3} = 0.0413 \text{ g PH}_3$$

$$0.0413 \text{ g PH}_3 \times \frac{1000 \text{ mg PH}_3}{1 \text{ g}} = 41.3 \text{ mg PH}_3$$

- (b) number of molecules of $\text{PH}_3 = 0.001216 \text{ mol PH}_3 \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol PH}_3}$
 number of molecules of $\text{PH}_3 = 7.32 \times 10^{20}$ molecules

- 21.** At the higher elevation of the mountains, the atmospheric pressure is lower than at the beach. However, the bag is virtually leak proof; no gas escapes. Thus, the gas inside the bag expands in the lower pressure until the bag is filled to near bursting. (It would have been difficult to predict this result. The temperature in the mountains is usually lower than at the beach. The lower temperature would *decrease* the pressure of the gas.)

General Gas Equation

- 23.** Because the number of moles of gas does not change, $\frac{P_i \times V_i}{T_i} = nR = \frac{P_f \times V_f}{T_f}$ is

obtained from the ideal gas equation. This expression can be rearranged as follows.

$$V_f = \frac{V_i \times P_i \times T_f}{P_f \times T_i} = \frac{4.25 \text{ L} \times 748 \text{ mmHg} \times (273.2 + 26.8)\text{K}}{742 \text{ mmHg} \times (273.2 + 25.6)\text{K}} = 4.30 \text{ L}$$

25. Volume and pressure are constant. Hence $n_i T_i = \frac{PV}{R} = n_f T_f$

$$\frac{n_f}{n_i} = \frac{T_i}{T_f} = \frac{(21 + 273.15) \text{ K}}{(210 + 273.15) \text{ K}} = 0.609 \text{ (60.9 \% of the gas remains)}$$

Hence, 39.1% of the gas must be released. Mass of gas released = $12.5 \text{ g} \times \frac{39.1}{100} = 4.89 \text{ g}$

Ideal Gas Equation

27. Assume that the $\text{CO}_2(\text{g})$ behaves ideally and use the ideal gas law: $PV = nRT$

$$V = \frac{nRT}{P} = \frac{\left(89.2 \text{ g} \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g}}\right) 0.08206 \frac{\text{L atm}}{\text{mol K}} (37 + 273.2) \text{ K}}{737 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 5.32 \times 10^4 \text{ mL}$$

29. $\text{mass} = n \times M = \frac{PV}{RT}$ $M = \frac{11.2 \text{ atm} \times 18.5 \text{ L} \times 83.80 \text{ g/mol}}{0.08206 \frac{\text{L atm}}{\text{mol K}} (28.2 + 273.2) \text{ K}} = 702 \text{ g Kr}$

31. $n_{\text{gas}} = 5.0 \times 10^9 \text{ molecules gas} \times \frac{1 \text{ mol gas}}{6.022 \times 10^{23} \text{ molecules gas}} = 8.3 \times 10^{-15} \text{ mol gas}$

We next determine the pressure that the gas exerts at 25 °C in a cubic meter

$$P = \frac{8.3 \times 10^{-15} \text{ mol gas} \times 0.08206 \frac{\text{L atm}}{\text{K mol}} \times 298.15 \text{ K}}{1 \text{ m}^3 \times \left(\frac{10 \text{ dm}}{1 \text{ m}}\right)^3 \times \left(\frac{1 \text{ L}}{1 \text{ dm}^3}\right)} \times \frac{101,325 \text{ Pa}}{1 \text{ atm}} = 2.1 \times 10^{-11} \text{ Pa}$$

33. The basic ideal law relationship applies here. Molar volume is the amount of volume that one mole of a gas occupies. If $PV = nRT$, then molar volume is V/n , and the relationship rearranges to:

$$\frac{V}{n} = V_m = \frac{R \cdot T}{P}$$

(a) $V_m = (298 \text{ K} \cdot 0.08206 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1}) / 1.00 \text{ atm} = 24.4 \text{ L} \cdot \text{mol}^{-1}$

(b) $P_{\text{atm}} = 748 \text{ mmHg} / 760 \text{ mmHg} = 0.978 \text{ atm}$

$V_m = (373 \text{ K} \cdot 0.08206 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1}) / 0.978 \text{ atm} = 31.3 \text{ L} \cdot \text{mol}^{-1}$

Determining Molar Mass

- 35.** Use the ideal gas law to determine the amount in moles of the given quantity of gas.

$$M = \frac{mRT}{PV} = \frac{0.418 \text{ g} \times 0.08206 \frac{\text{L atm}}{\text{mol K}} \times 339.5 \text{ K}}{\left(743 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}\right) \times 0.115 \text{ L}} = 104 \text{ g mol}^{-1}$$

Alternatively

$$n = \frac{PV}{RT} = \frac{\left(743 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}\right) \left(115 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}}\right)}{0.08206 \frac{\text{L atm}}{\text{mol K}} (273.2 + 66.3) \text{ K}} = 0.00404 \text{ mol gas} \quad M = \frac{0.418 \text{ g}}{0.00404 \text{ mol}} = 103 \text{ g/mol}$$

- 37.** First we determine the empirical formula for the sulfur fluoride. Assume a 100 g sample of S_xF_y .

$$\text{moles S} = 29.6 \text{ g S} \times \frac{1 \text{ mol S}}{32.064 \text{ g S}} = 0.923 \text{ mol S} \quad \text{moles F} = 70.4 \text{ g F} \times \frac{1 \text{ mol F}}{18.9984 \text{ g F}} = 3.706 \text{ mol F}$$

Dividing the number of moles of each element by 0.923 moles gives the empirical formula SF_4 .

To find the molar mass we use the relationship:

$$\text{Molar mass} = \frac{dRT}{P} = \frac{4.5 \text{ g L}^{-1} \times 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 293 \text{ K}}{1.0 \text{ atm}} = 108 \text{ g mol}^{-1}$$

$$\text{Thus, molecular formula} = (\text{empirical formula}) \times \left(\frac{\text{molecular formula mass}}{\text{empirical formula mass}}\right) = \frac{108 \text{ g mol}^{-1}}{108.06 \text{ g mol}^{-1}} \times SF_4 = SF_4$$

39. (a)
$$M = \frac{mRT}{PV} = \frac{0.231 \text{ g} \times 0.08206 \frac{\text{L atm}}{\text{mol K}} \times (23 + 273) \text{ K}}{\left(749 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}\right) \times \left(102 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}}\right)} = 55.8 \text{ g/mol}$$

- (b) The formula contains 4 atoms of carbon. (5 atoms of carbon gives a molar mass of at least 60—too high—and 3 C atoms gives a molar mass of 36—too low to be made up by adding H's.) To produce a molar mass of 56 with 4 carbons requires the inclusion of 8 atoms of H in the formula of the compound. Thus the formula is C_4H_8 .

Gas Densities

41.
$$d = \frac{MP}{RT} \quad \rightarrow \quad P = \frac{dRT}{M} = \frac{1.80 \text{ g/L} \times 0.08206 \frac{\text{L atm}}{\text{mol K}} \times (32 + 273) \text{ K}}{28.0 \text{ g/mol}} \times \frac{760 \text{ mmHg}}{1 \text{ atm}}$$

$$P = 1.21 \times 10^3 \text{ mmHg}$$

$$\text{Molar volume} = \frac{1 \text{ L}}{1.8 \text{ g}} N_2 \times \frac{28.0 \text{ g N}_2}{1 \text{ mol N}_2} = 15.56 \text{ L/mol}$$

$$\mathbf{43. (a)} \quad d = \frac{MP}{RT} = \frac{28.96 \text{ g/mol} \times 1.00 \text{ atm}}{0.08206 \frac{\text{L atm}}{\text{mol K}} \times (273 + 25)\text{K}} = 1.18 \text{ g/L air}$$

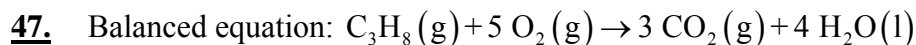
$$\mathbf{(b)} \quad d = \frac{MP}{RT} = \frac{44.0 \text{ g/mol CO}_2 \times 1.00 \text{ atm}}{0.08206 \frac{\text{L atm}}{\text{mol K}} \times (273 + 25)\text{K}} = 1.80 \text{ g/L CO}_2$$

Since this density is greater than that of air, the balloon will not rise in air when filled with CO₂ at 25°C; instead, it will sink!

$$\mathbf{45.} \quad d = \frac{MP}{RT} \text{ becomes } M = \frac{dRT}{P} = \frac{2.64 \text{ g/L} \times 0.08206 \frac{\text{L atm}}{\text{mol K}} \times (310+273)\text{K}}{775 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}} = 124 \text{ g/mol}$$

Since the atomic mass of phosphorus is 31.0, the formula of phosphorus molecules in the vapor must be P₄. (4 atoms/molecule × 31.0 = 124)

Gases in Chemical Reactions



Use the law of combining volumes. O₂ volume = 75.6 L C₃H₈ × $\frac{5 \text{ L O}_2}{1 \text{ L C}_3\text{H}_8}$ = 378 L O₂

49. Determine the moles of SO₂(g) produced and then use the ideal gas equation.

Conversion pathway approach:

$$\text{mol SO}_2 = \left(1.2 \times 10^6 \text{ kg coal} \times \frac{3.28 \text{ kg S}}{100.00 \text{ kg coal}} \times \frac{1000 \text{ g S}}{1 \text{ kg S}} \times \frac{1 \text{ mol S}}{32.1 \text{ g S}} \times \frac{1 \text{ mol SO}_2}{1 \text{ mol S}} \right)$$

$$= 1.23 \times 10^6 \text{ mol SO}_2$$

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

$$V = \frac{1.23 \times 10^6 \text{ mol SO}_2 \times 0.08206 \frac{\text{L atm}}{\text{mol K}} \times 296 \text{ K}}{738 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}} = 3.1 \times 10^7 \text{ L SO}_2$$

Stepwise approach:

$$1.2 \times 10^6 \text{ kg coal} \times \frac{3.28 \text{ kg S}}{100.00 \text{ kg coal}} = 3.94 \times 10^4 \text{ kg S}$$

$$3.94 \times 10^4 \text{ kg S} \times \frac{1000 \text{ g S}}{1 \text{ kg S}} = 3.94 \times 10^7 \text{ g S}$$

$$3.94 \times 10^7 \text{ g S} \times \frac{1 \text{ mol S}}{32.1 \text{ g S}} = 1.23 \times 10^6 \text{ mol S}$$

$$1.23 \times 10^6 \text{ mol S} \times \frac{1 \text{ mol SO}_2}{1 \text{ mol S}} = 1.23 \times 10^6 \text{ mol SO}_2$$

$$V = \frac{nRT}{P} = \frac{1.23 \times 10^6 \text{ mol SO}_2 \times 0.08206 \frac{\text{L atm}}{\text{mol K}} \times 296 \text{ K}}{738 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}} = 3.1 \times 10^7 \text{ L SO}_2$$

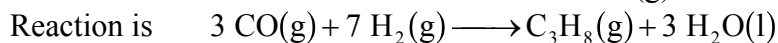
51. Determine the moles of O₂, and then the mass of KClO₃ that produced this amount of O₂.

$$\text{mol O}_2 = \frac{\left(738 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}\right) \times \left(119 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}}\right)}{0.08206 \frac{\text{L atm}}{\text{mol K}} \times (22.4 + 273.2) \text{ K}} = 0.00476 \text{ mol O}_2$$

$$\text{mass KClO}_3 = 0.00476 \text{ mol O}_2 \times \frac{2 \text{ mol KClO}_3}{3 \text{ mol O}_2} \times \frac{122.6 \text{ g KClO}_3}{1 \text{ mol KClO}_3} = 0.389 \text{ g KClO}_3$$

$$\% \text{ KClO}_3 = \frac{0.389 \text{ g KClO}_3}{3.57 \text{ g sample}} \times 100\% = 10.9\% \text{ KClO}_3$$

53. First we need to find the number of moles of CO(g)



$$n_{\text{CO}} = \frac{PV}{RT} = \frac{28.5 \text{ L} \times 760 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}}{0.08206 \frac{\text{L atm}}{\text{K mol}} \times 273.15 \text{ K}} = 1.27 \text{ moles CO}$$

$$V_{\text{H}_2(\text{required})} = \frac{n_{\text{H}_2} RT}{P} = \frac{1.27 \text{ mol CO} \times \frac{7 \text{ mol H}_2}{3 \text{ mol CO}} \times 0.08206 \frac{\text{L atm}}{\text{K mol}} \times 299 \text{ K}}{751 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}} = 73.7 \text{ LH}_2$$

Mixtures of Gases

- 55.** Determine the total amount of gas; then use the ideal gas law, assuming that the gases behave ideally.

$$\begin{aligned} \text{moles gas} &= \left(15.2 \text{ g Ne} \times \frac{1 \text{ mol Ne}}{20.18 \text{ g Ne}} \right) + \left(34.8 \text{ g Ar} \times \frac{1 \text{ mol Ar}}{39.95 \text{ g Ar}} \right) \\ &= 0.753 \text{ mol Ne} + 0.871 \text{ mol Ar} = 1.624 \text{ mol gas} \\ V &= \frac{nRT}{P} = \frac{1.624 \text{ mol} \times 0.08206 \frac{\text{L atm}}{\text{mol K}} \times (26.7 + 273.2) \text{ K}}{7.15 \text{ atm}} = 5.59 \text{ L gas} \end{aligned}$$

- 57.** The two pressures are related, as are the number of moles of $\text{N}_2(\text{g})$ to the total number of moles of gas.

$$\begin{aligned} \text{moles N}_2 &= \frac{PV}{RT} = \frac{28.2 \text{ atm} \times 53.7 \text{ L}}{0.08206 \frac{\text{L atm}}{\text{mol K}} \times (26 + 273) \text{ K}} = 61.7 \text{ mol N}_2 \\ \text{total moles of gas} &= 61.7 \text{ mol N}_2 \times \frac{75.0 \text{ atm}}{28.2 \text{ atm}} = 164 \text{ mol gas} \\ \text{mass Ne} &= (164 \text{ mol total} - 61.7 \text{ mol N}_2) \times \frac{20.18 \text{ g Ne}}{1 \text{ mol Ne}} = 2.06 \times 10^3 \text{ g Ne} \end{aligned}$$

- 59.** Initial pressure of the cylinder

$$P = \frac{nRT}{V} = \frac{(1.60 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{31.998 \text{ g O}_2})(0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1})(273.15 \text{ K})}{2.24 \text{ L}} = 0.500 \text{ atm}$$

We need to quadruple the pressure from 0.500 atm to 2.00 atm.

The mass of O_2 needs to quadruple as well from 1.60 g \rightarrow 6.40 g or add 4.80 g O_2

(this answer eliminates answer (a) and (b) as being correct).

One could also increase the pressure by adding the same number of another gas (e.g. He)

$$\text{mass of He} = n_{\text{He}} \times \text{MM}_{\text{He}}$$

$$\text{(Note: moles of O}_2 \text{ needed} = 4.80 \text{ g} \times \frac{1 \text{ mol O}_2}{31.998 \text{ g O}_2} = 0.150 \text{ moles} = 0.150 \text{ moles of He)}$$

$$\text{mass of He} = 0.150 \text{ moles} \times \frac{4.0026 \text{ g He}}{1 \text{ mol He}} = 0.600 \text{ g He} \text{ ((d) is correct, add 0.600 g of He)}$$

61.

$$\begin{aligned}
 \text{(a)} \quad P_{\text{ben}} &= \frac{nRT}{V} = \frac{\left(0.728 \text{ g} \times \frac{1 \text{ mol C}_6\text{H}_6}{78.11 \text{ g C}_6\text{H}_6}\right) \times 0.08206 \frac{\text{L atm}}{\text{mol K}} \times (35 + 273) \text{ K}}{2.00 \text{ L}} \times \frac{760 \text{ mmHg}}{1 \text{ atm}} \\
 &= 89.5 \text{ mmHg} \\
 P_{\text{total}} &= 89.5 \text{ mmHg C}_6\text{H}_6(\text{g}) + 752 \text{ mmHg Ar}(\text{g}) = 842 \text{ mmHg}
 \end{aligned}$$

$$\text{(b)} \quad P_{\text{benzene}} = 89.5 \text{ mmHg} \quad P_{\text{Ar}} = 752 \text{ mmHg}$$

63. 1.00 g H₂ ≈ 0.50 mol H₂ 1.00 g He ≈ 0.25 mol He

Adding 1.00 g of He to a vessel that only contains 1.00 g of H₂ results in the number of moles of gas being increased by 50%. Situation (b) best represents the resulting mixture, as the volume has increased by 50%

65. In this problem, you don't need to explicitly solve for moles of gas, since you are looking at the relationship between pressure and volume.

$$\text{mol O}_2 = \frac{PV}{RT} = \frac{(4.0 \text{ atm})(1.0 \text{ L})}{RT} = \frac{4.0}{RT}$$

$$\text{mol N}_2 = \frac{PV}{RT} = \frac{(2.0 \text{ atm})(2.0 \text{ L})}{RT} = \frac{4.0}{RT}$$

$$\text{total mol. of gas} = 8.0/RT$$

Therefore,

$$P = \frac{nRT}{V} = \frac{8.0}{RT} \cdot \frac{RT}{2.0} = 4.0 \text{ atm}$$

Collecting Gases over Liquids

67. The pressure of the liberated H₂(g) is 744 mmHg – 23.8 mmHg = 720. mmHg

$$V = \frac{nRT}{P} = \frac{\left(1.65 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g}} \times \frac{3 \text{ mol H}_2}{2 \text{ mol Al}}\right) 0.08206 \frac{\text{L atm}}{\text{mol K}} (273 + 25) \text{ K}}{720. \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}} = 2.37 \text{ L H}_2(\text{g})$$

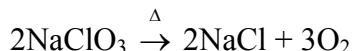
This is the total volume of both gases, each with a different partial pressure.

- 69.** We first determine the pressure of the gas collected. This would be its “dry gas” pressure and, when added to 22.4 mmHg, gives the barometric pressure.

$$P = \frac{nRT}{V} = \frac{\left(1.46 \text{ g} \times \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2}\right) 0.08206 \frac{\text{L atm}}{\text{mol K}} \times 297 \text{ K}}{1.16 \text{ L}} \times \frac{760 \text{ mmHg}}{1 \text{ atm}} = 729 \text{ mmHg}$$

barometric pressure = 729 mm Hg + 22.4 mmHg = 751 mmHg

- 71.** The first step is to balance the equation:



The pressure of O₂ is determined by subtracting the known vapor pressure of water at the given temperature from the measured total pressure.

$$P_{\text{O}_2} = P_{\text{TOT}} - P_{\text{H}_2\text{O}} = 734 \text{ torr} - 21.07 \text{ torr} = 713 \text{ torr}$$

$$P_{\text{atm}} = 713 \text{ mmHg} / 760 \text{ mmHg} = 0.938 \text{ atm}$$

$$\text{mol O}_2 = \frac{PV}{RT} = \frac{(0.938 \text{ atm})(0.0572 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1})(296 \text{ K})} = 0.00221 \text{ mol}$$

Mass of NaClO₃ is then determined as follows:

$$0.00221 \text{ mol O}_2 \times \frac{2 \text{ mol NaClO}_3}{3 \text{ mol O}_2} \times \frac{106.44 \text{ g}}{1 \text{ mol NaClO}_3} = 0.1568 \text{ g NaClO}_3$$

$$\% \text{NaClO}_3 = \frac{0.1568 \text{ g}}{0.8765 \text{ g}} \times 100 = 17.9\%$$

Kinetic-Molecular Theory

- 73.** Recall that 1 J = kg m² s⁻²

$$u_{\text{rms}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3 \times 8.3145 \frac{\text{J}}{\text{mol K}} \times \frac{1 \text{ kg m}^2 \text{ s}^{-2}}{1 \text{ J}} \times 303 \text{ K}}{\frac{70.91 \times 10^{-3} \text{ kg Cl}_2}{1 \text{ mol Cl}_2}}} = 326 \text{ m/s}$$

- 75.**

$$M = \frac{3RT}{(u_{\text{rms}})^2} = \frac{3 \times 8.3145 \frac{\text{J}}{\text{mol K}} \times 298 \text{ K}}{\left(2180 \frac{\text{mi}}{\text{hr}} \times \frac{1 \text{ hr}}{3600 \text{ sec}} \times \frac{5280 \text{ ft}}{1 \text{ mi}} \times \frac{12 \text{ in.}}{1 \text{ ft}} \times \frac{1 \text{ m}}{39.37 \text{ in.}}\right)^2} = 0.00783 \text{ kg/mol} = 7.83 \text{ g/mol. or } 7.83 u.$$

- 77.** We equate the two expressions for root mean square speed, cancel the common factors, and solve for the temperature of Ne. Note that the units of molar masses do not have to be in kg/mol in this calculation; they simply must be expressed in the same units.

$$\sqrt{\frac{3RT}{M}} = \sqrt{\frac{3R \times 300 \text{ K}}{4.003}} = \sqrt{\frac{3R \times T_{\text{Ne}}}{20.18}} \quad \text{Square both sides:} \quad \frac{300 \text{ K}}{4.003} = \frac{T_{\text{Ne}}}{20.18}$$

$$\text{Solve for } T_{\text{Ne}}: \quad T_{\text{Ne}} = 300 \text{ K} \times \frac{20.18}{4.003} = 1.51 \times 10^3 \text{ K}$$

- 79.** The greatest pitfall of this type of problem is using improper units. Therefore, convert everything to SI units.

$$\text{MM O}_2 = 32.0 \text{ g/mol} = 32.0 \times 10^{-3} \text{ kg/mol}$$

$$\text{Mass of O}_2 \text{ molecule: } \frac{32.0 \times 10^{-3} \text{ kg}}{\text{mol}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molec.}} = 5.314 \times 10^{-26} \text{ kg}$$

$$R = 8.3145 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}, \text{ or } \text{kg} \cdot \text{m}^2 / (\text{s}^2 \cdot \text{mol} \cdot \text{K})$$

Now, we must determine the u_{rms} first to determine kinetic energy:

$$\overline{u_{rms}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3 \times 8.3145 \times 298}{32.0 \times 10^{-3}}} = 482 \text{ m/s}$$

Kinetic energy of an O₂ molecule is as follows:

$$\overline{e_k} = \frac{1}{2} m \overline{u_{rms}^2} = \frac{1}{2} (5.314 \times 10^{-26} \text{ kg})(482 \text{ m/s})^2 = 6.17 \times 10^{-21} \text{ J/molecule}$$

Diffusion and Effusion of Gases

81.

$$\frac{\text{rate (NO}_2\text{)}}{\text{rate (N}_2\text{O)}} = \sqrt{\frac{M(\text{N}_2\text{O})}{M(\text{NO}_2)}} = \sqrt{\frac{44.02}{46.01}} = 0.9781 = \frac{x \text{ mol NO}_2/t}{0.00484 \text{ mol N}_2\text{O}/t}$$

$$\text{mol NO}_2 = 0.00484 \text{ mol} \times 0.9781 = 0.00473 \text{ mol NO}_2$$

83.

$$\begin{array}{ll} \text{(a)} \quad \frac{\text{rate (N}_2\text{)}}{\text{rate (O}_2\text{)}} = \sqrt{\frac{M(\text{O}_2)}{M(\text{N}_2)}} = \sqrt{\frac{32.00}{28.01}} = 1.07 & \text{(b)} \quad \frac{\text{rate (H}_2\text{O)}}{\text{rate (D}_2\text{O)}} = \sqrt{\frac{M(\text{D}_2\text{O})}{M(\text{H}_2\text{O})}} = \sqrt{\frac{20.0}{18.02}} = 1.05 \\ \text{(c)} \quad \frac{\text{rate (}^{14}\text{CO}_2\text{)}}{\text{rate (}^{12}\text{CO}_2\text{)}} = \sqrt{\frac{M(^{12}\text{CO}_2)}{M(^{14}\text{CO}_2)}} = \sqrt{\frac{44.0}{46.0}} = 0.978 & \text{(d)} \quad \frac{\text{rate (}^{235}\text{UF}_6\text{)}}{\text{rate (}^{238}\text{UF}_6\text{)}} = \sqrt{\frac{M(^{238}\text{UF}_6)}{M(^{235}\text{UF}_6)}} = \sqrt{\frac{352}{349}} = 1.004 \end{array}$$

- 85.** For ideal gases, the effusion rate is inversely proportional to their molecular mass. As such, the rate of effusion of a known gas can be determined if the rate of effusion for another gas is known:

$$\frac{\text{rate of effusion of Ne}}{\text{rate of effusion of He}} = \sqrt{\frac{MM_{\text{He}}}{MM_{\text{Ne}}}}$$

Since effusion is loosely defined as movement of a fixed number of atoms per unit time, and since in this problem we are looking at the time it takes for the same number of moles of both Ne and He to effuse, the above equation can be rearranged as follows:

$$\frac{\text{mol}_{\text{Ne}} / \text{time}_{\text{Ne}}}{\text{mol}_{\text{He}} / \text{time}_{\text{He}}} = \frac{\text{time}_{\text{He}}}{\text{time}_{\text{Ne}}} = \sqrt{\frac{MM_{\text{He}}}{MM_{\text{Ne}}}}$$

$$\frac{x}{22 \text{ h}} = \sqrt{\frac{4.00}{20.18}}$$

Solving for x ,

$$x = 22 \times \sqrt{4.00 / 20.18} = 9.80 \text{ h}$$

Nonideal Gases

- 87.** For $\text{Cl}_2(\text{g})$, $n^2a = 6.49 \text{ L}^2 \text{ atm}$ and $nb = 0.0562 \text{ L}$. $P_{\text{vdw}} = \frac{nRT}{V - nb} - \frac{n^2a}{V^2}$

At 0°C , $P_{\text{vdw}} = 9.90 \text{ atm}$ and $P_{\text{ideal}} = 11.2 \text{ atm}$, off by 1.3 atm or +13%

$$\text{(a) At } 100^\circ\text{C } P_{\text{ideal}} = \frac{nRT}{V} = \frac{1.00 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{mol K}} \times 373 \text{ K}}{2.00 \text{ L}} = 15.3 \text{ atm}$$

$$P_{\text{vdw}} = \frac{1.00 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{mol K}} \times T}{(2.00 - 0.0562) \text{ L}} - \frac{6.49 \text{ L}^2 \text{ atm}}{(2.00 \text{ L})^2} = 0.0422T \text{ atm} - 1.62 \text{ atm}$$

$$= 0.0422 \times 373 \text{ K} - 1.62 = 14.1 \text{ atm} \quad P_{\text{ideal}} \text{ is off by } 1.2 \text{ atm or } +8.5\%$$

$$\text{(b) At } 200^\circ\text{C } P_{\text{ideal}} = \frac{nRT}{V} = \frac{1.00 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{mol K}} \times 473 \text{ K}}{2.00 \text{ L}} = 19.4 \text{ atm}$$

$$P_{\text{vdw}} = 0.0422_2 \times 473 \text{ K} - 1.62_3 = 18.3_5 \text{ atm} \quad P_{\text{ideal}} \text{ is off by } 1.0 \text{ atm or } +5.5\%$$

$$\text{(c) At } 400^\circ\text{C } P_{\text{ideal}} = \frac{nRT}{V} = \frac{1.00 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{mol K}} \times 673 \text{ K}}{2.00 \text{ L}} = 27.6 \text{ atm}$$

$$P_{\text{vdw}} = 0.0422 \times 673 \text{ K} - 1.62 = 26.8 \text{ atm} \quad P_{\text{ideal}} \text{ is off by } 0.8 \text{ atm or } +3.0\%$$

- 89.** The van der Waals parameter b is defined as the excluded volume per mole, or the volume that is taken up by 1 mole of gas once converted to a liquid.

From Table 6-5, $b_{\text{He}} = 0.0238 \text{ L/mol}$. Therefore, the volume of a single He atom is:

Conversion pathway approach:

$$\frac{0.0238 \text{ L}}{\text{mol He}} \times \frac{1 \text{ mol He}}{6.022 \times 10^{23} \text{ atoms}} \times \frac{1 \text{ m}^3}{1000 \text{ L}} \times \frac{(1 \times 10^{12} \text{ pm})^3}{1 \text{ m}^3} = 3.95 \times 10^7 \text{ pm}^3 / \text{He atom}$$

Step-wise approach:

$$\frac{0.0238 \text{ L}}{\text{mol He}} \times \frac{1 \text{ mol He}}{6.022 \times 10^{23} \text{ atoms}} = 3.95 \times 10^{-26} \text{ L/atom}$$

$$3.95 \times 10^{-26} \text{ L/atom} \times \frac{1 \text{ m}^3}{1000 \text{ L}} = 3.95 \times 10^{-29} \text{ m}^3/\text{atom}$$

$$3.95 \times 10^{-29} \text{ m}^3/\text{atom} \times \frac{(1 \times 10^{12} \text{ pm})^3}{1 \text{ m}^3} = 3.95 \times 10^7 \text{ pm}^3 / \text{He atom}$$

$V = (4/3) \pi r^3$. Rearranging to solve for r gives $r = \sqrt[3]{(3V)/(4\pi)}$. Solving for r gives an atomic radius of 211.3 pm.

Integrative and Advanced Exercises

- 94.** We know that the sum of the moles of gas in the two bulbs is 1.00 moles, and that both bulbs have the same volume, and are at the same pressure because they are connected. Therefore,

$$n_1 T_1 = n_2 T_2$$

$$\frac{n_1}{n_2} = \frac{T_2}{T_1} = \frac{350 \text{ K}}{225 \text{ K}} = 1.556$$

$$n_1 = 1.556 n_2$$

Therefore,

$n_2(1.556) + n_2 = 1.00$. Solving the equation gives (Flask 2) $n_2 = 0.391$, and (Flask 1) $n_1 = 0.609$.

97. *Stepwise approach:*

Note that three moles of gas are produced for each mole of NH_4NO_3 that decomposes.

$$\text{amount of gas} = 3.05 \text{ g NH}_4\text{NO}_3 \times \frac{1 \text{ mol NH}_4\text{NO}_3}{80.04 \text{ g NH}_4\text{NO}_3} \times \frac{3 \text{ moles of gas}}{1 \text{ mol NH}_4\text{NO}_3} = 0.114 \text{ mol gas}$$

$$T = 250 + 273 = 523 \text{ K}$$

$$P = \frac{nRT}{V} = \frac{0.114 \text{ mol} \times 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 523 \text{ K}}{2.18 \text{ L}} = 2.25 \text{ atm}$$

Conversion pathway:

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

$$P = \frac{3.05 \text{ g NH}_4\text{NO}_3 \times \frac{1 \text{ mol NH}_4\text{NO}_3}{80.04 \text{ g NH}_4\text{NO}_3} \times \frac{3 \text{ moles of gas}}{1 \text{ mol NH}_4\text{NO}_3} \times 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 523 \text{ K}}{2.18 \text{ L}}$$

$$= 2.25 \text{ atm}$$

98. First, let's convert the given units to those easier used:

$$P = 101 \text{ kPa} \times (1 \text{ barr}/101 \text{ kPa}) \times (1 \text{ atm}/1.01 \text{ barr}) = 0.9901 \text{ atm}$$

$$T = 819 \text{ }^\circ\text{C} + 273 \text{ K} = 1092 \text{ K}$$

$$\text{mol NH}_4\text{NO}_2 = 128 \text{ g NH}_4\text{NO}_2 \times \frac{1 \text{ mol NH}_4\text{NO}_2}{64.052 \text{ g}} = 1.998 \text{ mol}$$

$$\text{mol gas} = 1.998 \text{ mol NH}_4\text{NO}_2 \times \frac{3 \text{ mol gas}}{1 \text{ mol NH}_4\text{NO}_2} = 5.994$$

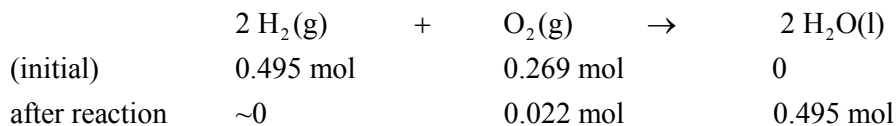
$$V = \frac{nRT}{P} = \frac{(5.994 \text{ mol})(0.082058 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1})(1092 \text{ K})}{0.9901 \text{ atm}} = 542 \text{ L}$$

99.

$$(a) n_{\text{H}_2} = 1.00 \text{ g}/2.02 \text{ g/mol} = 0.495 \text{ mol H}_2; n_{\text{O}_2} = 8.60 \text{ g}/32.0 \text{ g/mol} = 0.269 \text{ mol O}_2$$

$$P_{\text{total}} = P_{\text{O}_2} + P_{\text{H}_2} = \frac{n_{\text{total}}RT}{V} = \frac{0.769 \text{ mol} \times 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 298 \text{ K}}{1.50 \text{ L}} = 12.5 \text{ atm}$$

The limiting reagent in the production of water is H_2 (0.495 mol) with O_2 (0.269 mol) in excess.



$$\text{so } P_{\text{total}} = P_{\text{O}_2} + P_{\text{H}_2\text{O}}$$

$$P_{\text{O}_2} = \frac{n_{\text{total}}RT}{V} = \frac{0.022 \text{ mol} \times 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 298 \text{ K}}{1.50 \text{ L}} = 0.359 \text{ atm}$$

$$P_{\text{total}} = 0.359 \text{ atm} + \frac{23.8 \text{ mm Hg}}{760 \text{ mm Hg/atm}} = 0.39 \text{ atm}$$

101. Determine relative numbers of moles, and the mole fractions of the 3 gases in the 100.0 g gaseous mixture.

$$\text{amount } N_2 = 46.5 \text{ g } N_2 \times \frac{1 \text{ mol } N_2}{28.01 \text{ g } N_2} = 1.66 \text{ mol } N_2 \div 2.86 \text{ mol} \longrightarrow 0.580$$

$$\text{amount } Ne = 12.7 \text{ g } Ne \times \frac{1 \text{ mol } Ne}{20.18 \text{ g } Ne} = 0.629 \text{ mol } Ne \div 2.86 \text{ mol} \longrightarrow 0.220$$

$$\text{amount } Cl_2 = 40.8 \text{ g } Cl_2 \times \frac{1 \text{ mol } Cl_2}{70.91 \text{ g } Cl_2} = 0.575 \text{ mol } Cl_2 \div 2.86 \text{ mol} \longrightarrow 0.201$$

$$\text{total amount} = 1.66 \text{ mol } N_2 + 0.629 \text{ mol } Ne + 0.575 \text{ mol } Cl_2 = 2.86 \text{ mol}$$

Since the total pressure of the mixture is 1 atm, the partial pressure of each gas is numerically very close to its mole fraction. Thus, the partial pressure of Cl_2 is 0.201 atm or 153 mmHg

102. Let us first determine the molar mass of this mixture.

$$M = \frac{dRT}{P} = \frac{0.518 \text{ g } L^{-1} \times 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 298 \text{ K}}{721 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}} = 13.4 \text{ g/mol}$$

Then we let x be the mole fraction of He in the mixture.

$$13.4 \text{ g/mol} = x \times 4.003 \text{ g/mol} + (1.000 - x) \times 32.00 \text{ g/mol} = 4.003x + 32.00 - 32.00x$$

$$x = \frac{32.00 - 13.4}{32.00 - 4.003} = 0.664$$

Thus, one mole of the mixture contains 0.664 mol He. We determine the mass of that He and then the % He by mass in the mixture.

$$\text{mass He} = 0.664 \text{ mol He} \times \frac{4.003 \text{ g He}}{1 \text{ mol He}} = 2.66 \text{ g He} \quad \% \text{He} = \frac{2.66 \text{ g He}}{13.4 \text{ g mixture}} \times 100\% = 19.9\% \text{ He}$$

105. The amount of $N_2(g)$ plus the amount of $He(g)$ equals the total amount of gas. We use this equality, and substitute with the ideal gas law, letting V symbolize the volume of cylinder B.

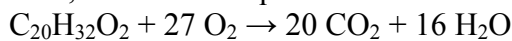
$$n_{N_2} + n_{He} = n_{\text{total}} \quad \text{becomes} \quad P_{N_2} \times V_{N_2} + P_{He} \times V_{He} = P_{\text{total}} \times V_{\text{total}}$$

since all temperatures are equal.

$$\text{Substitution gives } (8.35 \text{ atm} \times 48.2 \text{ L}) + (9.50 \text{ atm} \times V) = 8.71 \text{ atm} (48.2 \text{ L} + V)$$

$$402 \text{ L atm} + 9.50V = 420 \text{ L atm} + 8.71V \quad V = \frac{420 - 402}{9.50 - 8.71} = 23 \text{ L}$$

108. First, balance the equation:



$$\begin{aligned} \text{mol O}_2 \text{ needed: } & 2000 \text{ g C}_{20}\text{H}_{32}\text{O}_2 \times \frac{1 \text{ mol C}_{20}\text{H}_{32}\text{O}_2}{304.52 \text{ g C}_{20}\text{H}_{32}\text{O}_2} \times \frac{27 \text{ mol O}_2}{1 \text{ mol C}_{20}\text{H}_{32}\text{O}_2} \\ & = 177.33 \text{ mol O}_2 \end{aligned}$$

Using the ideal gas law, we can determine the volume of 177.33 mol of O₂:

$$\text{vol O}_2 = \frac{nRT}{P} = \frac{(177.33 \text{ mol})(0.082058 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1})(298 \text{ K})}{1.00 \text{ atm}} = 4336.30 \text{ L O}_2$$

To determine the volume of air needed, we note that O₂ represents 20.9% of air by volume: $x(0.205) = 4336.30 \text{ L}$. Solving for x gives 20698 L, or $2.070 \times 10^4 \text{ L}$.

109. First recognize that 3 moles of gas are produced from every 2 moles of water, and compute the number of moles of gas produced. Then determine the partial pressure these gases would exert.

$$\text{amount of gas} = 1.32 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{3 \text{ mol gas}}{2 \text{ mol H}_2\text{O}} = 0.110 \text{ mol gas}$$

$$P = \frac{nRT}{V} = \frac{0.110 \text{ mol} \times 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 303 \text{ K}}{2.90 \text{ L}} \times \frac{760 \text{ mmHg}}{1 \text{ atm}} = 717 \text{ mmHg}$$

Then the vapor pressure (partial pressure) of water is determined by difference.

$$P_{\text{water}} = 748 \text{ mmHg} - 717 \text{ mmHg} = 31 \text{ mmHg}$$

111. The total pressure of the mixture of O₂ and H₂O is 737 mmHg, and the partial pressure of H₂O is 25.2 mmHg.

(a) The percent of water vapor by volume equals its percent pressure.

$$\% \text{ H}_2\text{O} = \frac{25.2 \text{ mmHg}}{737 \text{ mmHg}} \times 100\% = 3.42\% \text{ H}_2\text{O by volume}$$

(b) The % water vapor by number of molecules equals its percent pressure, 3.42% by number.

(c) One mole of the combined gases contains 0.0342 mol H₂O and 0.9658 mol O₂.

$$\begin{aligned} \text{molar mass} &= 0.0342 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} + 0.9658 \text{ mol O}_2 \times \frac{31.999 \text{ g O}_2}{1 \text{ mol O}_2} \\ &= 0.616 \text{ g H}_2\text{O} + 30.90 \text{ g O}_2 = 31.52 \text{ g} \end{aligned}$$

$$\% \text{ H}_2\text{O} = \frac{0.616 \text{ g H}_2\text{O}}{31.52 \text{ g}} \times 100\% = 1.95\% \text{ H}_2\text{O by mass}$$

- 112.** (a) 1 mol of the mixture at STP occupies a volume of 22.414 L. It contains 0.79 mol He and 0.21 mol O₂.

$$\text{STP density} = \frac{\text{mass}}{\text{volume}} = \frac{0.79 \text{ mol He} \times \frac{4.003 \text{ g He}}{1 \text{ mol He}} + 0.21 \text{ mol O}_2 \times \frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2}}{22.414 \text{ L}} = 0.44 \text{ g/L}$$

25°C is a temperature higher than STP. This condition increases the 1.00-L volume containing 0.44 g of the mixture at STP. We calculate the expanded volume with the combined gas law.

$$V_{\text{final}} = 1.00 \text{ L} \times \frac{(25 + 273.2) \text{ K}}{273.2 \text{ K}} = 1.09 \text{ L} \quad \text{final density} = \frac{0.44 \text{ g}}{1.09 \text{ L}} = 0.40 \text{ g/L}$$

We determine the apparent molar masses of each mixture by multiplying the mole fraction (numerically equal to the volume fraction) of each gas by its molar mass, and then summing these products for all gases in the mixture.

$$M_{\text{air}} = \left(0.78084 \times \frac{28.01 \text{ g N}_2}{1 \text{ mol N}_2} \right) + \left(0.20946 \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) + \left(0.00934 \times \frac{39.95 \text{ g Ar}}{1 \text{ mol Ar}} \right)$$

$$= 21.87 \text{ g N}_2 + 6.703 \text{ g O}_2 + 0.373 \text{ g Ar} = 28.95 \text{ g/mol air}$$

$$M_{\text{mix}} = \left(0.79 \times \frac{4.003 \text{ g He}}{1 \text{ mol He}} \right) + \left(0.21 \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 3.2 \text{ g He} + 6.7 \text{ g O}_2 = \frac{9.9 \text{ g mixture}}{\text{mol}}$$

- (b) In order to prepare two gases with the same density, the volume of the gas of smaller molar mass must be smaller by a factor equal to the ratio of the molar masses. According to Boyle's law, this means that the pressure on the less dense gas must be larger by a factor equal to a ratio of molar masses.

$$P_{\text{mix}} = \frac{28.95}{9.9} \times 1.00 \text{ atm} = 2.9 \text{ atm}$$

- 114.** First, determine the moles of Na₂S₂O₃, then use the chemical equations given to determine the moles of O₃ in the mixture:

$$\text{mol Na}_2\text{S}_2\text{O}_3 = 0.0262 \text{ L} \times 0.1359 \text{ M} = 0.003561 \text{ mol}$$

$$\text{mol O}_3: 0.003561 \text{ mol Na}_2\text{S}_2\text{O}_3 \times \frac{1 \text{ mol I}_3^-}{2 \text{ mol Na}_2\text{S}_2\text{O}_3} \times \frac{1 \text{ mol O}_3}{1 \text{ mol I}_3^-} = 0.001780 \text{ mol O}_3$$

Using the ideal gas law, we can determine the moles of gas:

$$\text{moles of gas} = \frac{(0.993 \text{ atm})(53.2 \text{ L})}{(0.082058 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1})(291 \text{ K})} = 2.2123 \text{ mol}$$

$$\chi_{\text{O}_3} = \frac{0.001780 \text{ mol}}{2.2123 \text{ mol total}} = 8.046 \times 10^{-4}$$

115. (D) First compute the pressure of the water vapor at 30.1°C.

$$P_{\text{H}_2\text{O}} = \frac{\left(0.1052 \text{ g} \times \frac{1 \text{ mol H}_2\text{O}}{18.015 \text{ g H}_2\text{O}}\right) \times \frac{0.08206 \text{ L atm}}{\text{mol K}} \times (30.1 + 273.2) \text{ K}}{8.050 \text{ L}} \times \frac{760 \text{ mmHg}}{1 \text{ atm}} = 13.72 \text{ mmHg}$$

The water vapor is kept in the 8.050-L container, which means that its pressure is proportional to the absolute temperature in the container. Thus, for each of the six temperatures, we need to calculate two numbers: (1) the pressure due to this water (because gas pressure varies with temperature), and (2) 80% of the vapor pressure. The temperature we are seeking is where the two numbers agree.

$$P_{\text{water}}(T) = 13.72 \text{ mmHg} \times \frac{(T + 273.2) \text{ K}}{(30.1 + 273.2) \text{ K}}$$

$$P(20^\circ\text{C}) = 13.72 \text{ mmHg} \times \frac{(20. + 273.2) \text{ K}}{(30.1 + 273.2) \text{ K}} = 13.3 \text{ mmHg}$$

For example,

T	20. °C	19. °C	18. °C	17. °C	16. °C	15. °C
P _{water} , mmHg	13.3	13.2	13.2	13.1	13.1	13.0
80.0% v.p., mmHg	14.0	13.2	12.4	11.6	10.9	10.2

At approximately 19°C, the relative humidity of the air will be 80.0%.

118. (a) $\overline{u_{\text{rms}}}$ is determined as follows:

$$\overline{u_{\text{rms}}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3(8.3145 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1})(298 \text{ K})}{32.00 \times 10^{-3} \text{ kg}}} = 482 \text{ m/s}$$

$$F_u = 4\pi \left(\frac{M}{2\pi RT}\right)^{3/2} u^2 \exp(-Mu^2 / 2RT)$$

$$(b) F_u = 4\pi \left(\frac{32.00 \times 10^{-3}}{2\pi(8.3145)(298)}\right)^{3/2} (498)^2 \exp\left(-\frac{(32.00 \times 10^{-3})(498)^2}{2(8.3145)(298)}\right) = 1.92 \times 10^{-3}$$

119. Potential energy of an object is highest when the kinetic energy of the object is zero and the object has attained its maximum height. Therefore, we must determine the kinetic energy. But first, we have to determine the velocity of the N₂ molecule.

$$\overline{u_{\text{rms}}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3(8.3145 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1})(300 \text{ K})}{28.00 \times 10^{-3} \text{ kg}}} = 517 \text{ m/s}$$

$$\overline{e_k} = \frac{1}{2} m \overline{u_{\text{rms}}}^2 = \frac{1}{2} (28.00 \times 10^{-3} \text{ kg})(517 \text{ m/s})^2 = 3742 \text{ J}$$

$$\overline{e_k} = \overline{e_p} \Rightarrow m \cdot g \cdot h = (28.00 \times 10^{-3} \text{ kg})(9.8 \text{ m/s}^2) \cdot h = 3742 \text{ J}$$

Solving for h, the altitude reached by an N₂ molecule is 13637 m or 13.6 km.

121. (a) First multiply out the left-hand side of the equation.

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT = PV - Pnb + \frac{an^2}{V} - \frac{abn^3}{V^2}$$

Now multiply the entire equation through by V^2 , and collect all terms on the right-hand side.

$$0 = -nRTV^2 + PV^3 - PnbV^2 + an^2V - abn^3$$

Finally, divide the entire equation by P and collect terms with the same power of V , to obtain:

$$0 = V^3 - n\left(\frac{RT + bP}{P}\right)V^2 + \left(\frac{n^2a}{P}\right)V - \frac{n^3ab}{P} = 0$$

(b)

$$n = 185 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.0 \text{ g CO}_2} = 4.20 \text{ mol CO}_2$$

$$\begin{aligned} 0 &= V^3 - 4.20 \text{ mol} \left(\frac{(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 286 \text{ K}) + (0.0429 \text{ L/mol} \times 12.5 \text{ atm})}{12.5 \text{ atm}} \right) V^2 \\ &\quad + \left(\frac{(4.20 \text{ mol})^2 \cdot 3.61 \text{ L}^2 \text{ atm mol}^{-2}}{12.5 \text{ atm}} \right) V - \frac{(4.20 \text{ mol})^3 \cdot 3.61 \text{ L}^2 \text{ atm mol}^{-2} \times 0.0429 \text{ L/mol}}{12.5 \text{ atm}} \\ &= V^3 - 8.07 V^2 + 5.09 V - 0.918 \end{aligned}$$

We can solve this equation by the method of successive approximations. As a first value, we use the volume obtained from the ideal gas equation:

$$V = \frac{nRT}{P} = \frac{4.20 \text{ mol} \times 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 286 \text{ K}}{12.5 \text{ atm}} = 7.89 \text{ L}$$

	$V = 7.89 \text{ L}$	$(7.89)^3 - 8.06(7.89)^2 + (5.07 \times 7.89) - 0.909 = 28.5$	>0
Try	$V = 7.00 \text{ L}$	$(7.00)^3 - 8.06(7.00)^2 + (5.07 \times 7.00) - 0.909 = -17.4$	<0
Try	$V = 7.40 \text{ L}$	$(7.40)^3 - 8.06(7.40)^2 + (5.07 \times 7.40) - 0.909 = 0.47$	>0
Try	$V = 7.38 \text{ L}$	$(7.38)^3 - 8.06(7.38)^2 + (5.07 \times 7.38) - 0.909 = -0.53$	<0
Try	$V = 7.39 \text{ L}$	$(7.39)^3 - 8.06(7.39)^2 + (5.07 \times 7.39) - 0.909 = -0.03$	<0

The volume of CO_2 is very close to 7.39 L.

A second way is to simply disregard the last term and to solve the equation

$$0 = V^3 - 8.06 V^2 + 5.07 V$$

This equation simplifies to the following quadratic equation:

$0 = V^2 - 8.06 V + 5.07$, which is solved with the quadratic formula.

$$V = \frac{8.06 \pm \sqrt{(8.06)^2 - 4 \times 5.07}}{2} = \frac{8.06 \pm 6.68}{2} = \frac{14.74}{2} = 7.37 \text{ L}$$

The other root, 0.69 L does not appear to be reasonable, due to its small size.

FEATURE PROBLEMS

- 126.** Nitryl Fluoride $65.01 \text{ u} \left(\frac{49.4}{100} \right) = 32.1 \text{ u of X}$
 Nitrosyl Fluoride $49.01 \text{ u} \left(\frac{32.7}{100} \right) = 16.0 \text{ u of X}$
 Thionyl Fluoride $86.07 \text{ u} \left(\frac{18.6}{100} \right) = 16.0 \text{ u of X}$
 Sulfuryl Fluoride $102.07 \text{ u} \left(\frac{31.4}{100} \right) = 32.0 \text{ u of X}$

The atomic mass of X is 16 u which corresponds to the element oxygen.

The number of atoms of X (oxygen) in each compound is given below:

Nitryl Fluoride = 2 atoms of O	Nitrosyl Fluoride = 1 atom of O
Thionyl Fluoride = 1 atom of O	Sulfuryl Fluoride = 2 atoms of O

- 127. (a)** The $\text{N}_2(\text{g})$ extracted from liquid air has some $\text{Ar}(\text{g})$ mixed in. Only $\text{O}_2(\text{g})$ was removed from liquid air in the oxygen-related experiments.
- (b)** Because of the presence of $\text{Ar}(\text{g})$ [39.95 g/mol], the $\text{N}_2(\text{g})$ [28.01 g/mol] from liquid air will have a greater density than $\text{N}_2(\text{g})$ from nitrogen compounds.

(c) Magnesium will react with molecular nitrogen $[3 \text{Mg}(\text{s}) + \text{N}_2(\text{g}) \rightarrow \text{Mg}_3\text{N}_2(\text{s})]$ but not with Ar. Thus, magnesium reacts with all the nitrogen in the mixture, but leaves the relatively inert $\text{Ar}(\text{g})$ unreacted.

(d) The “nitrogen” remaining after oxygen is extracted from each mole of air (Rayleigh’s mixture) contains $0.78084 + 0.00934 = 0.79018 \text{ mol}$ and has the mass calculated below.

$$\text{mass of gaseous mixture} = (0.78084 \times 28.013 \text{ g/mol N}_2) + (0.00934 \times 39.948 \text{ g/mol Ar})$$

$$\text{mass of gaseous mixture} = 21.874 \text{ g N}_2 + 0.373 \text{ g Ar} = 22.247 \text{ g mixture.}$$

Then, the molar mass of the mixture can be computed: $22.247 \text{ g mixture} / 0.79018 \text{ mol} = 28.154 \text{ g/mol}$. Since the STP molar volume of an ideal gas is 22.414 L, we can compute the two densities.

$$d(\text{N}_2) = \frac{28.013 \text{ g/mol}}{22.414 \text{ L/mol}} = 1.2498 \text{ g/mol} \quad d(\text{mixture}) = \frac{28.154 \text{ g/mol}}{22.414 \text{ L/mol}} = 1.2561 \text{ g/mol}$$

These densities differ by 0.50%.

129. Total mass = mass of payload + mass of balloon + mass of H₂

Use ideal gas law to calculate mass of H₂: $PV = nRT = \frac{\text{mass}(m)}{\text{Molar mass}(M)} RT$

$$m = \frac{PVM}{RT} = \frac{1 \text{ atm} \left(120 \text{ ft}^3 \times \frac{(12 \text{ in})^3}{(1 \text{ ft})^3} \times \frac{(2.54 \text{ cm})^3}{(1 \text{ in})^3} \times \frac{1 \times 10^{-3} \text{ L}}{1 \text{ cm}^3} \right) 2.016 \frac{\text{g}}{\text{mol}}}{\left(0.08206 \frac{\text{L atm}}{\text{K mol}} \right) 273 \text{ K}} = 306 \text{ g}$$

$$\text{Total mass} = 1200 \text{ g} + 1700 \text{ g} + 306 \text{ g} \approx 3200 \text{ g}$$

We know at the maximum height, the balloon will be 25 ft in diameter. Need to find out what mass of air is displaced. We need to make one assumption – the volume percent of air is unchanged with altitude. Hence we use an apparent molar mass for air of 29 g mol⁻¹ (question 99). Using the data provided, we find the altitude at which the balloon displaces 3200 g of air.

$$\text{Note: balloon radius} = \frac{25}{2} = 12.5 \text{ ft. volume} = \frac{4}{3} (\pi)r^3 = \frac{4}{3} (3.1416)(12.5)^3 = 8181 \text{ ft}^3$$

$$\text{Convert to liters: } 8181 \text{ ft}^3 \times \frac{(12 \text{ in})^3}{(1 \text{ ft})^3} \times \frac{(2.54 \text{ cm})^3}{(1 \text{ in})^3} \times \frac{1 \times 10^{-3} \text{ L}}{1 \text{ cm}^3} = 231,660 \text{ L}$$

$$\text{At 10 km: } m = \frac{PVM}{RT} = \frac{\left(2.7 \times 10^2 \text{ mb} \times \frac{1 \text{ atm}}{1013.25 \text{ mb}} \right) (231,660 \text{ L}) \left(29 \frac{\text{g}}{\text{mol}} \right)}{\left(0.08206 \frac{\text{L atm}}{\text{K mol}} \right) 223 \text{ K}} = 97,827 \text{ g}$$

$$\text{At 20 km: } m = \frac{PVM}{RT} = \frac{\left(5.5 \times 10^1 \text{ mb} \times \frac{1 \text{ atm}}{1013.25 \text{ mb}} \right) (231,660 \text{ L}) \left(29 \frac{\text{g}}{\text{mol}} \right)}{\left(0.08206 \frac{\text{L atm}}{\text{K mol}} \right) 217 \text{ K}} = 20,478 \text{ g}$$

$$\text{At 30 km: } m = \frac{PVM}{RT} = \frac{\left(1.2 \times 10^1 \text{ mb} \times \frac{1 \text{ atm}}{1013.25 \text{ mb}} \right) (231,660 \text{ L}) \left(29 \frac{\text{g}}{\text{mol}} \right)}{\left(0.08206 \frac{\text{L atm}}{\text{K mol}} \right) 230 \text{ K}} = 4,215 \text{ g}$$

$$\text{At 40 km: } m = \frac{PVM}{RT} = \frac{\left(2.9 \times 10^0 \text{ mb} \times \frac{1 \text{ atm}}{1013.25 \text{ mb}} \right) (231,660 \text{ L}) \left(29 \frac{\text{g}}{\text{mol}} \right)}{\left(0.08206 \frac{\text{L atm}}{\text{K mol}} \right) 250 \text{ K}} = 937 \text{ g}$$

The lifting power of the balloon will allow it to rise to an altitude of just over 30 km.

SELF-ASSESSMENT EXERCISES

133. The answer is (d). The following shows the values for each:

(a) $P = g \cdot h \cdot d = (9.8 \text{ m/s}^2)(0.75 \text{ m})(13600 \text{ kg/m}^3) = 99960 \text{ Pa}$

(b) Just for a rough approximation, we assume the density of air to be the same as that of nitrogen (this underestimates it a bit, but is close enough). The density of N_2 is $0.02802 \text{ g/22.7 L} = 1.234 \text{ g/L} = 1.234 \text{ kg/m}^3$. $P = (9.8 \text{ m/s}^2)(16093 \text{ m})(1.234 \text{ kg/m}^3) = 194616 \text{ Pa}$.

(c) $P = g \cdot h \cdot d = (9.8 \text{ m/s}^2)(5.0 \text{ m})(1590 \text{ kg/m}^3) = 77910 \text{ Pa}$

(d) $P = nRT/V = (10.00 \text{ g H}_2 \times 1 \text{ mol/2.02 g})(0.083145 \text{ L}\cdot\text{barr}\cdot\text{K}^{-1})(273 \text{ K})/(22.7 \text{ L}) = 4.95 \text{ barr} = 495032 \text{ Pa}$

134. The answer is (c), because the temperature decreases from $100 \text{ }^\circ\text{C}$ (373 K) to 200 K.

135. $P_1/T_1 = P_2/T_2$. To calculate T_2 , rearrange the formula:

$$T_2 = 2.0 \text{ barr} / (1.0 \text{ barr} \times 273 \text{ K}) = 546 \text{ K}$$

136. The answer is (d).

$$\frac{P_1 \cdot V_1}{T_1} = \frac{P_2 \cdot V_2}{T_2}$$

$$V_2 = \frac{1 \text{ atm} \cdot 22.4 \text{ L} \cdot 298 \text{ K}}{273 \text{ K} \cdot 1.5 \text{ atm}} = 16.3 \text{ L}$$

137. The answer is (b). Since the same number of moles of all ideal gases occupy the same volume, density is driven by the molar mass of the gas. Therefore, Kr has the highest density because it has a molar mass of 83 g/mol.

138. The answer is (a). Using the formula $u_{\text{rms}} = (3RT/M)^{1/2}$, increases T by a factor of 2 increase u_{rms} by a factor of $\sqrt{2}$.

139. (a) False. They both have the same kinetic energy

(b) True. All else being equal, the heavier molecule is slower.

(c) False. The formula $PV=nRT$ can be used to confirm this. The answer is 24.4 L

(d) True. There is ~ 1.0 mole of each gas present. All else being equal, the same number of moles of any ideal gas occupies the same volume.

(e) False. Total pressure is the sum of the partial pressures. So long as there is nothing else but H_2 and O_2 , the total pressure is equal to the sum of the individual partial pressures.

140. The answer is (c). Partial pressures are additive, so:

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

$$P_{\text{O}_2} (\text{atm}) = \frac{P_{\text{TOT}} - P_{\text{H}_2\text{O}}}{760 \text{ mm Hg}} = \frac{751 - 21}{760} = 0.96 \text{ atm}$$

141. The answer is (a). First, determine the # moles of NH_3 using the ideal gas law relationship:
 $n = PV/RT = (0.500 \text{ atm} \times 4.48 \text{ L}) / (0.08206 \text{ L}\cdot\text{atm K}^{-1} \times 273 \text{ K}) = 0.100 \text{ mol}$
 If 1.0 mole of a substance has 6.022×10^{23} atoms, 0.100 moles has 6.022×10^{22} atoms.

142. The answer is (b). Since $PV = nRT$, the number of moles of O_2 needed to satisfy the conditions of the problem is:

$$n = (2.00 \text{ atm} \times 2.24 \text{ L}) / (0.08206 \text{ L}\cdot\text{atm K}^{-1} \times 273 \text{ K}) = 0.200 \text{ moles}$$

$$\text{The amount of } \text{O}_2 \text{ available} = 1.60 \text{ g } \text{O}_2 \times \frac{1 \text{ mol } \text{O}_2}{32.0 \text{ g } \text{O}_2} = 0.050 \text{ mol } \text{O}_2$$

To have an additional 0.150 mol in the system, we would add 0.6 g of He to the container.

143. The volumes of both gases were measured at the same temperature and pressure. Therefore, the proportionality constant between volume and moles for both gases is the same (that is, volume can essentially replace moles in the following calculations):

$$25.0 \text{ L } \text{H}_2 \times \frac{3 \text{ L CO}}{7 \text{ L H}_2} = 10.7 \text{ L CO needed}$$

So, all of the H_2 and 10.7 L of CO are consumed, and 1.3 L of CO remain.

145. The answer is (c). Gases behave more ideally at high temperatures and low pressures.

- 146.** (a) He or Ne: Ne has higher a and b values.
 (b) CH_4 or C_3H_8 : C_3H_8 has higher a and b values
 (c) H_2 or Cl_2 : Cl_2 has higher a and b values

147. We know that pressure is force per unit area, that is:

$$P = \frac{F}{A} = \frac{m \cdot g}{A}$$

Using the fact that area $A = \pi r^2 = \pi(D/2)^2$ and that mass $m = d \cdot V$, and that volume of a cylindrical tube is $V = A \cdot h$ (where h is the height of the liquid in the tube), we can express the pressure formula as follows:

$$P = \frac{F}{A} = \frac{d \cdot V \cdot g}{A} = \frac{d \cdot (A \cdot h) \cdot g}{A} = d \cdot h \cdot g$$

Therefore,

$$\frac{d \cdot V \cdot g}{A} = d \cdot h \cdot g, \text{ and,}$$

$$h = \frac{d}{A} = \frac{d}{\pi(D/2)^2}$$

As we can see, the height, h is inversely proportional to D . That is, the larger the diameter of the tube, the shorter the height of the liquid.

- 148.** First, convert the given information to more useful units. The pressure (752 torr) is equivalent to 0.9895 atm (752 mm/760 mm Hg), and the temperature is 298 K. Then, use the ideal gas relationship to determine how many moles of gas are present, assuming 1 L of gas:

$$PV = nRT$$

$$n = \frac{PV}{RT} = \frac{(0.9895 \text{ atm})(1 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm K}^{-1})(298 \text{ K})} = 0.04046 \text{ mol}$$

From the problem, we know the mass of the gas per 1 L, as expressed in the density. Mass of 1 L of this gas is 2.35 g. The molar mass of this substance is therefore:

$$\text{MM} = \text{g/mol} = 2.35 \text{ g}/0.04046 \text{ mol} = 58.082 \text{ g/mol}$$

Now, let's calculate the empirical formula for the hydrocarbon gas and see how it compares to the molar mass:

$$\text{mol C} = 82.7 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 6.89 \text{ mol}$$

$$\text{mol H} = 17.3 \text{ g H} \times \frac{1 \text{ mol H}}{1.01 \text{ g H}} = 17.1 \text{ mol}$$

$$\text{H:C ratio} = 17.1 \text{ mol} / 6.89 \text{ mol} = 2.5$$

The ratio between H and C is 2.5:1 or 5:2, making the empirical formula C_2H_5 . The mass of this formula unit is 29.07 g/mol. Comparing to the calculated molar mass of the gas, it is smaller by a factor of 2. Therefore, the gas in question is C_4H_{10} or butane.

- 149.** N_2 comprises 78.084% of atmosphere, oxygen 20.946%, argon 0.934%, and CO_2 0.0379%. To graphically show the scale of this difference, divide all values by the smallest one (CO_2). Therefore, for every single mark representing CO_2 , we need 2060 marks for N_2 , 553 marks for O_2 , and 25 for Ar.