

## CHAPTER 5

### GASES

#### Pressure

21. a.  $4.8 \text{ atm} \times \frac{760 \text{ mm Hg}}{\text{atm}} = 3.6 \times 10^3 \text{ mm Hg}$ ; b.  $3.6 \times 10^3 \text{ mm Hg} \times \frac{1 \text{ torr}}{\text{mm Hg}}$   
 $= 3.6 \times 10^3 \text{ torr}$

c.  $4.8 \text{ atm} \times \frac{1.013 \times 10^5 \text{ Pa}}{\text{atm}} = 4.9 \times 10^5 \text{ Pa}$ ; d.  $4.8 \text{ atm} \times \frac{14.7 \text{ psi}}{\text{atm}} = 71 \text{ psi}$

22. If the levels of mercury in each arm of the manometer are equal, then the pressure in the flask is equal to atmospheric pressure. When they are unequal, the difference in height in millimeters will be equal to the difference in pressure in millimeters of mercury between the flask and the atmosphere. Which level is higher will tell us whether the pressure in the flask is less than or greater than atmospheric.

a.  $P_{\text{flask}} < P_{\text{atm}}$ ;  $P_{\text{flask}} = 760. - 118 = 642 \text{ torr}$

$$642 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.845 \text{ atm}$$

$$0.845 \text{ atm} \times \frac{1.013 \times 10^5 \text{ Pa}}{\text{atm}} = 8.56 \times 10^4 \text{ Pa}$$

b.  $P_{\text{flask}} > P_{\text{atm}}$ ;  $P_{\text{flask}} = 760. \text{ torr} + 215 \text{ torr} = 975 \text{ torr}$

$$975 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.28 \text{ atm}$$

$$1.28 \text{ atm} \times \frac{1.013 \times 10^5 \text{ Pa}}{\text{atm}} = 1.30 \times 10^5 \text{ Pa}$$

c.  $P_{\text{flask}} = 635 - 118 = 517 \text{ torr}$ ;  $P_{\text{flask}} = 635 + 215 = 850. \text{ torr}$

$$23. \quad 4.75 \text{ cm} \times \frac{10 \text{ mm}}{\text{cm}} = 47.5 \text{ mm Hg or } 47.5 \text{ torr}; \quad 47.5 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 6.25 \times 10^{-2} \text{ atm}$$

$$6.25 \times 10^{-2} \text{ atm} \times \frac{1.013 \times 10^5 \text{ Pa}}{\text{atm}} = 6.33 \times 10^3 \text{ Pa}$$

24. a. The pressure is proportional to the mass of the fluid. The mass is proportional to the volume of the column of fluid (or to the height of the column, assuming the area of the column of fluid is constant).

$d = \text{density} = \frac{\text{mass}}{\text{volume}}$ ; the volume of silicon oil is the same as the volume of mercury in Exercise 22.

$$V = \frac{m}{d}; \quad V_{\text{Hg}} = V_{\text{oil}}; \quad \frac{m_{\text{Hg}}}{d_{\text{Hg}}} = \frac{m_{\text{oil}}}{d_{\text{oil}}}, \quad m_{\text{oil}} = \frac{m_{\text{Hg}} d_{\text{oil}}}{d_{\text{Hg}}}$$

Because P is proportional to the mass of liquid:

$$P_{\text{oil}} = P_{\text{Hg}} \left( \frac{d_{\text{oil}}}{d_{\text{Hg}}} \right) = P_{\text{Hg}} \left( \frac{1.30}{13.6} \right) = (0.0956)P_{\text{Hg}}$$

This conversion applies only to the column of silicon oil.

a.  $P_{\text{flask}} = 760. \text{ torr} - (118 \times 0.0956) \text{ torr} = 760. - 11.3 = 749 \text{ torr}$

$$749 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.986 \text{ atm}; \quad 0.986 \text{ atm} \times \frac{1.013 \times 10^5 \text{ Pa}}{\text{atm}} = 9.99 \times 10^4 \text{ Pa}$$

b.  $P_{\text{flask}} = 760. \text{ torr} + (215 \times 0.0956) \text{ torr} = 760. + 20.6 = 781 \text{ torr}$

$$781 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.03 \text{ atm}; \quad 1.03 \text{ atm} \times \frac{1.013 \times 10^5 \text{ Pa}}{\text{atm}} = 1.04 \times 10^5 \text{ Pa}$$

- b. If we are measuring the same pressure, the height of the silicon oil column would be  $13.6/1.30 = 10.5$  times the height of a mercury column. The advantage of using a less dense fluid than mercury is in measuring small pressures. The height difference measured will be larger for the less dense fluid. Thus the measurement will be more precise.
25. Suppose we have a column of mercury  $1.00 \text{ cm} \times 1.00 \text{ cm} \times 76.0 \text{ cm} = V = 76.0 \text{ cm}^3$ :

$$\text{mass} = 76.0 \text{ cm}^3 \times 13.59 \text{ g/cm}^3 = 1.03 \times 10^3 \text{ g} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 1.03 \text{ kg}$$

$$F = mg = 1.03 \text{ kg} \times 9.81 \text{ m/s}^2 = 10.1 \text{ kg m/s}^2 = 10.1 \text{ N}$$

$$\frac{\text{Force}}{\text{Area}} = \frac{10.1 \text{ N}}{\text{cm}^2} \times \left(\frac{100 \text{ cm}}{\text{m}}\right)^2 = 1.01 \times 10^5 \frac{\text{N}}{\text{m}^2} \text{ or } 1.01 \times 10^5 \text{ Pa}$$

(Note: 76.0 cm Hg = 1 atm =  $1.01 \times 10^5$  Pa.)

To exert the same pressure, a column of water will have to contain the same mass as the 76.0-cm column of mercury. Thus the column of water will have to be 13.59 times taller or  $76.0 \text{ cm} \times 13.59 = 1.03 \times 10^3 \text{ cm} = 10.3 \text{ m}$ .

### Gas Laws

$$26. \quad \frac{PV}{T} = nR = \text{constant}, \quad \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$P_2 = \frac{P_1 V_1 T_2}{V_2 T_1} = 710 \text{ torr} \times \frac{5.0 \times 10^2 \text{ mL}}{25 \text{ mL}} \times \frac{(273+820) \text{ K}}{(273+30.) \text{ K}} = 5.1 \times 10^4 \text{ torr}$$

$$27. \quad \frac{PV}{nT} = R; \text{ for a gas at two conditions:}$$

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}; \text{ because } n \text{ and } V \text{ are constant: } \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$T_2 = \frac{P_2 T_1}{P_1} = \frac{2500 \text{ torr} \times 294.2 \text{ K}}{758 \text{ torr}} = 970 \text{ K} = 7.0 \times 10^2 \text{ }^\circ\text{C}$$

For two-condition problems, units for P and V just need to be the same units for both conditions, not necessarily atm and L. The unit conversions from other P or V units would cancel when applied to both conditions. However, temperature always must be converted to the Kelvin scale. The temperature conversions between other units and Kelvin will not cancel each other.

$$28. \quad PV = nRT, \text{ } n \text{ is constant. } \frac{PV}{T} = nR = \text{constant}, \quad \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}, \quad V_2 = \frac{V_1 P_1 T_2}{P_2 T_1}$$

$$V_2 = 1.00 \text{ L} \times \frac{760. \text{ torr}}{220. \text{ torr}} \times \frac{(273-31) \text{ K}}{(273+23) \text{ K}} = 2.82 \text{ L}; \quad \Delta V = 2.82 - 1.00 = 1.82 \text{ L}$$

29. Treat each gas separately, and use the relationship  $P_1V_1 = P_2V_2$  ( $n$  and  $T$  are constant).

$$\text{For H}_2: P_2 = \frac{P_1V_1}{V_2} = 475 \text{ torr} \times \frac{2.00 \text{ L}}{3.00 \text{ L}} = 317 \text{ torr}$$

$$\text{For N}_2: P_2 = 0.200 \text{ atm} \times \frac{1.00 \text{ L}}{3.00 \text{ L}} = 0.0667 \text{ atm}; 0.0667 \text{ atm} \times \frac{760 \text{ torr}}{\text{atm}} = 50.7 \text{ torr}$$

$$P_{\text{total}} = P_{\text{H}_2} + P_{\text{N}_2} = 317 + 50.7 = 368 \text{ torr}$$

30. For H<sub>2</sub>:  $P_2 = \frac{P_1V_1}{V_2} = 360. \text{ torr} \times \frac{2.00 \text{ L}}{3.00 \text{ L}} = 240. \text{ torr}$

$$P_{\text{total}} = P_{\text{H}_2} + P_{\text{N}_2}, \quad P_{\text{N}_2} = P_{\text{total}} - P_{\text{H}_2} = 320. \text{ torr} - 240. \text{ torr} = 80. \text{ torr}$$

$$\text{For N}_2: P_1 = \frac{P_2V_2}{V_1} = 80. \text{ torr} \times \frac{3.00 \text{ L}}{1.00 \text{ L}} = 240 \text{ torr}$$

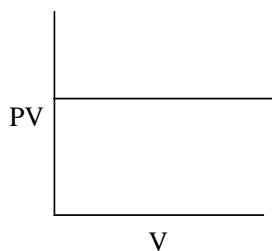
31. As NO<sub>2</sub> is converted completely into N<sub>2</sub>O<sub>4</sub>, the moles of gas present will decrease by a factor of one-half (from the 2 : 1 mol ratio in the balanced equation). Using Avogadro's law:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}, \quad V_2 = V_1 \times \frac{n_2}{n_1} = 25.0 \text{ mL} \times \frac{1}{2} = 12.5 \text{ mL}$$

N<sub>2</sub>O<sub>4</sub>(g) will occupy one-half the original volume of NO<sub>2</sub>(g).

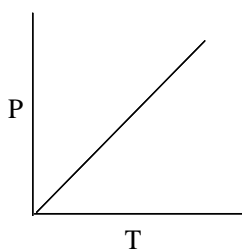
32. a.  $PV = nRT$

$$PV = \text{constant}$$



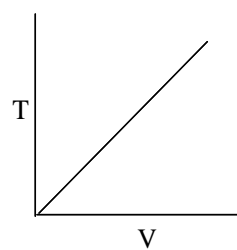
- b.  $PV = nRT$

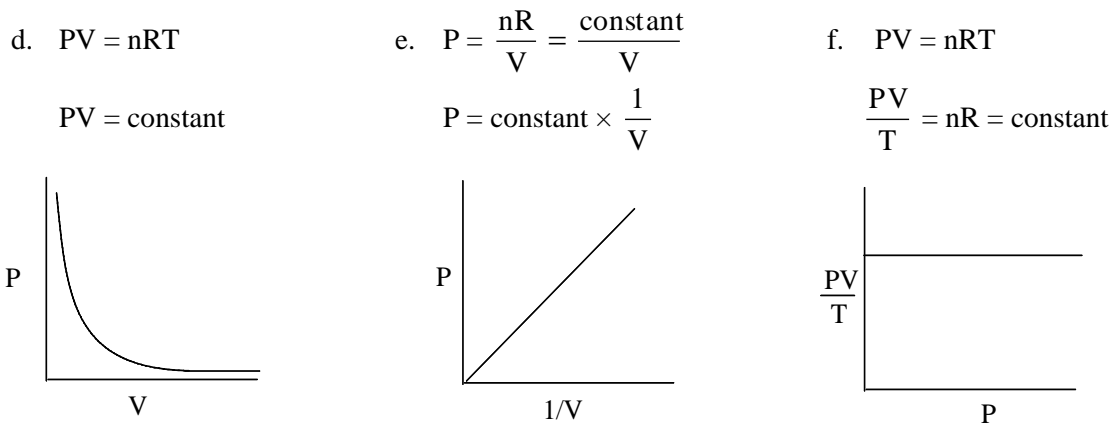
$$P = \left( \frac{nR}{V} \right) \times T = \text{const} \times T$$



- c.  $PV = nRT$

$$T = \left( \frac{P}{nR} \right) \times V = \text{const} \times V$$





*Note:* The equation for a straight line is  $y = mx + b$  where  $y$  is the  $y$  axis and  $x$  is the  $x$  axis. Any equation that has this form will produce a straight line with slope equal to  $m$  and a  $y$  intercept equal to  $b$ . Plots b, c, and e have this straight-line form.

33.  $PV = nRT$ ,  $P$  is constant.  $\frac{nT}{V} = \frac{P}{R} = \text{constant}$ ,  $\frac{n_1 T_1}{V_1} = \frac{n_2 T_2}{V_2}$

$$\frac{n_2}{n_1} = \frac{T_1 V_2}{T_2 V_1} = \frac{294 \text{ K}}{335 \text{ K}} \times \frac{4.20 \times 10^3 \text{ m}^3}{4.00 \times 10^3 \text{ m}^3} = 0.921$$

34. Because the container is flexible,  $P$  is assumed constant. The moles of gas present are also constant.

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}, \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}; \quad V_{\text{sphere}} = 4/3 \pi r^3$$

$$V_2 = \frac{V_1 T_2}{T_1}, \quad 4/3 \pi (r_2)^3 = \frac{4/3 \pi (1.00 \text{ cm})^3 \times 361 \text{ K}}{280. \text{ K}}$$

$$r_2^3 = \frac{361 \text{ K}}{280. \text{ K}} = 1.29, \quad r_2 = (1.29)^{1/3} = 1.09 \text{ cm} = \text{radius of sphere after heating}$$

35. The decrease in temperature causes the balloon to contract ( $V$  and  $T$  are directly related). Because weather balloons do expand, the effect of the decrease in pressure must be dominant.

36. For a gas at two conditions:  $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$

Because  $V$  is constant:  $\frac{P_1}{n_1 T_1} = \frac{P_2}{n_2 T_2}$ ,  $n_2 = \frac{n_1 P_2 T_1}{P_1 T_2}$

$$n_2 = \frac{1.50 \text{ mol} \times 800. \text{ torr} \times 298 \text{ K}}{400. \text{ torr} \times 323 \text{ K}} = 2.77 \text{ mol}$$

Moles of gas added =  $n_2 - n_1 = 2.77 - 1.50 = 1.27 \text{ mol}$

$$37. \quad n = \frac{PV}{RT} = \frac{135 \text{ atm} \times 200.0 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times (273 + 24) \text{ K}} = 1.11 \times 10^3 \text{ mol}$$

$$\text{For He: } 1.11 \times 10^3 \text{ mol} \times \frac{4.003 \text{ g He}}{\text{mol}} = 4.44 \times 10^3 \text{ g He}$$

$$\text{For H}_2: 1.11 \times 10^3 \text{ mol} \times \frac{2.016 \text{ g H}_2}{\text{mol}} = 2.24 \times 10^3 \text{ g H}_2$$

38. For the first diagram, there is a total volume of  $3X$  after the stopcock is open. The six total gas particles will be equally distributed (on average) over the entire volume ( $3X$ ). So per  $X$  volume, there will be two gas particles. Your first drawing should have four gas particles in the  $2X$  volume flask and two gas particles in the  $X$  volume flask.

Applying Boyle's law, the pressure in the two flasks after the stopcock is opened is:

$$P_1 V_1 = P_2 V_2, \quad P_2 = \frac{P_1 V_1}{V_2} = \frac{P_1 \times 2X}{3X} = \frac{2}{3} P_1$$

The final pressure in both flasks will be two-thirds that of the initial pressure in the left flask.

For the second diagram, there is a total volume of  $2X$  after the stopcock is opened. The gas particles will be equally distributed (on average) so that your drawing should have three gas particles in each flask. The final pressure is:

$$P_2 = \frac{P_1 V_1}{V_2} = \frac{P_1 \times X}{2X} = \frac{P_1}{2}$$

The final pressure in both flasks will be one-half that of the initial pressure in the left flask.

$$39. \quad P = P_{\text{CO}_2} = \frac{n_{\text{CO}_2} RT}{V} = \frac{\left(22.0 \text{ g} \times \frac{1 \text{ mol}}{44.01 \text{ g}}\right) \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 300. \text{ K}}{4.00 \text{ L}} = 3.08 \text{ atm}$$

With air present, the partial pressure of  $\text{CO}_2$  will still be  $3.08 \text{ atm}$ . The total pressure will be the sum of the partial pressures.

$$P_{\text{total}} = P_{\text{CO}_2} + P_{\text{air}} = 3.08 \text{ atm} + \left(740. \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}\right) = 3.08 + 0.974 = 4.05 \text{ atm}$$

$$40. \quad PV = nRT, \quad n \text{ is constant. } \frac{PV}{T} = nR = \text{constant, } \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$V_2 = (1.040)V_1, \quad \frac{V_1}{V_2} = \frac{1.000}{1.040}$$

$$P_2 = \frac{P_1 V_1 T_2}{V_2 T_1} = 75 \text{ psi} \times \frac{1.000}{1.040} \times \frac{(273+58) \text{ K}}{(273+19) \text{ K}} = 82 \text{ psi}$$

41.  $PV = nRT$ ,  $\frac{nT}{P} = \frac{V}{R} = \text{constant}$ ,  $\frac{n_1 T_1}{P_1} = \frac{n_2 T_2}{P_2}$ ; moles  $\times$  molar mass = mass

$$\frac{n_1(\text{molar mass})T_1}{P_1} = \frac{n_2(\text{molar mass})T_2}{P_2}, \quad \frac{\text{mass}_1 \times T_1}{P_1} = \frac{\text{mass}_2 \times T_2}{P_2}$$

$$\text{mass}_2 = \frac{\text{mass}_1 \times T_1 P_2}{T_2 P_1} = \frac{1.00 \times 10^3 \text{ g} \times 291 \text{ K} \times 650. \text{ psi}}{299 \text{ K} \times 2050. \text{ psi}} = 309 \text{ g}$$

42. If we had 100.0 g of the gas, we would have 50.0 g He and 50.0 g Xe.

$$\chi_{\text{He}} = \frac{n_{\text{He}}}{n_{\text{He}} + n_{\text{Xe}}} = \frac{\frac{50.0 \text{ g}}{4.003 \text{ g/mol}}}{\frac{50.0 \text{ g}}{4.003 \text{ g/mol}} + \frac{50.0 \text{ g}}{131.3 \text{ g/mol}}} = \frac{12.5 \text{ mol He}}{12.5 \text{ mol He} + 0.381 \text{ mol Xe}} = 0.970$$

$$P_{\text{He}} = \chi_{\text{He}} P_{\text{total}} = 0.970 \times 600. \text{ torr} = 582 \text{ torr}; \quad P_{\text{Xe}} = 600. - 582 = 18 \text{ torr}$$

43.  $P_{\text{He}} + P_{\text{H}_2\text{O}} = 1.00 \text{ atm} = 760. \text{ torr} = P_{\text{He}} + 23.8 \text{ torr}$ ,  $P_{\text{He}} = 736 \text{ torr}$

$$n_{\text{He}} = 0.586 \text{ g} \times \frac{1 \text{ mol}}{4.003 \text{ g}} = 0.146 \text{ mol He}$$

$$V = \frac{n_{\text{He}} RT}{P_{\text{He}}} = \frac{0.146 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298 \text{ K}}{736 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}} = 3.69 \text{ L}$$

44. The container has 5 He atoms, 3 Ne atoms, and 2 Ar atoms for a total of 10 atoms. The mole fractions of the various gases will be equal to the molecule fractions.

$$\chi_{\text{He}} = \frac{5 \text{ He atoms}}{10 \text{ total atoms}} = 0.50; \quad \chi_{\text{Ne}} = \frac{3 \text{ Ne atoms}}{10 \text{ total atoms}} = 0.30$$

$$\chi_{\text{Ar}} = 1.00 - 0.50 - 0.30 = 0.20$$

$$P_{\text{He}} = \chi_{\text{He}} \times P_{\text{total}} = 0.50(1.00 \text{ atm}) = 0.50 \text{ atm}$$

$$P_{\text{Ne}} = \chi_{\text{Ne}} \times P_{\text{Total}} = 0.30(1.00 \text{ atm}) = 0.30 \text{ atm}$$

$$P_{\text{Ar}} = 1.00 \text{ atm} - 0.50 \text{ atm} - 0.30 \text{ atm} = 0.20 \text{ atm}$$

$$45. \quad a. \quad \text{Mole fraction CH}_4 = \chi_{\text{CH}_4} = \frac{P_{\text{CH}_4}}{P_{\text{total}}} = \frac{0.175 \text{ atm}}{0.175 \text{ atm} + 0.250 \text{ atm}} = 0.412$$

$$\chi_{\text{O}_2} = 1.000 - 0.412 = 0.588$$

$$b. \quad PV = nRT, \quad n_{\text{total}} = \frac{P_{\text{total}} \times V}{RT} = \frac{0.425 \text{ atm} \times 10.5 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 338 \text{ K}} = 0.161 \text{ mol}$$

$$c. \quad \chi_{\text{CH}_4} = \frac{n_{\text{CH}_4}}{n_{\text{total}}}, \quad n_{\text{CH}_4} = \chi_{\text{CH}_4} \times n_{\text{total}} = 0.412 \times 0.161 \text{ mol} = 6.63 \times 10^{-2} \text{ mol CH}_4$$

$$6.63 \times 10^{-2} \text{ mol CH}_4 \times \frac{16.04 \text{ g CH}_4}{\text{mol CH}_4} = 1.06 \text{ g CH}_4$$

$$n_{\text{O}_2} = 0.588 \times 0.161 \text{ mol} = 9.47 \times 10^{-2} \text{ mol O}_2; \quad 9.47 \times 10^{-2} \text{ mol O}_2 \times \frac{32.00 \text{ g O}_2}{\text{mol O}_2} = 3.03 \text{ g O}_2$$

$$46. \quad P_{\text{total}} = 1.00 \text{ atm} = 760. \text{ torr} = P_{\text{N}_2} + P_{\text{H}_2\text{O}} = P_{\text{N}_2} + 17.5 \text{ torr}, \quad P_{\text{N}_2} = 743 \text{ torr}$$

$$n_{\text{N}_2} = \frac{P_{\text{N}_2} \times V}{RT} = \frac{(743 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}) \times (2.50 \times 10^2 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}})}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 293 \text{ K}} = 1.02 \times 10^{-2} \text{ mol N}_2$$

$$1.02 \times 10^{-2} \text{ mol N}_2 \times \frac{28.02 \text{ g N}_2}{\text{mol N}_2} = 0.286 \text{ g N}_2$$

47. We can use the ideal gas law to calculate the partial pressure of each gas or to calculate the total pressure. There will be less math if we calculate the total pressure from the ideal gas law.

$$n_{\text{O}_2} = 1.5 \times 10^2 \text{ mg O}_2 \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 4.7 \times 10^{-3} \text{ mol O}_2$$

$$n_{\text{NH}_3} = 5.0 \times 10^{21} \text{ molecules NH}_3 \times \frac{1 \text{ mol NH}_3}{6.022 \times 10^{23} \text{ molecules NH}_3} = 8.3 \times 10^{-3} \text{ mol NH}_3$$

$$n_{\text{total}} = n_{\text{N}_2} + n_{\text{O}_2} + n_{\text{NH}_3} = 5.0 \times 10^{-2} + 4.7 \times 10^{-3} + 8.3 \times 10^{-3} = 6.3 \times 10^{-2} \text{ mol total}$$

$$P_{\text{total}} = \frac{n_{\text{total}} \times RT}{V} = \frac{6.3 \times 10^{-2} \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 273 \text{ K}}{1.0 \text{ L}} = 1.4 \text{ atm}$$



$$P_{N_2} = \chi_{N_2} \times P_{\text{total}}, \quad \chi_{N_2} = \frac{n_{N_2}}{n_{\text{total}}}; \quad P_{N_2} = \frac{5.0 \times 10^{-2} \text{ mol}}{6.3 \times 10^{-2} \text{ mol}} \times 1.4 \text{ atm} = 1.1 \text{ atm}$$

$$P_{O_2} = \frac{4.7 \times 10^{-3}}{6.3 \times 10^{-2}} \times 1.4 \text{ atm} = 0.10 \text{ atm}; \quad P_{NH_3} = \frac{8.3 \times 10^{-3}}{6.3 \times 10^{-2}} \times 1.4 \text{ atm} = 0.18 \text{ atm}$$

48. For  $O_2$ ,  $n$  and  $T$  are constant, so  $P_1V_1 = P_2V_2$ .

$$P_1 = \frac{P_2V_2}{V_1} = 785 \text{ torr} \times \frac{1.94 \text{ L}}{2.00 \text{ L}} = 761 \text{ torr} = P_{O_2}$$

$$P_{\text{total}} = P_{O_2} + P_{H_2O}, \quad P_{H_2O} = 785 - 761 = 24 \text{ torr}$$

49. a. There are 6 He atoms and 4 Ne atoms, and each flask has the same volume. The He flask has 1.5 times as many atoms of gas present as the Ne flask, so the pressure in the He flask will be 1.5 times greater (assuming a constant temperature).
- b. Because the flask volumes are the same, your drawing should have the various atoms equally distributed between the two flasks. So each flask should have 3 He atoms and 2 Ne atoms.
- c. After the stopcock is opened, each flask will have 5 total atoms and the pressures will be equal. If six atoms of He gave an initial pressure of  $P_{\text{He, initial}}$ , then 5 total atoms will have a pressure of  $5/6 \times P_{\text{He, initial}}$ .

Using similar reasoning, 4 atoms of Ne gave an initial pressure of  $P_{\text{Ne, initial}}$ , so 5 total atoms will have a pressure of  $5/4 \times P_{\text{Ne, initial}}$ . Summarizing:

$$P_{\text{final}} = \frac{5}{6} P_{\text{He, initial}} = \frac{5}{4} P_{\text{Ne, initial}}$$

- d. For the partial pressures, treat each gas separately. For helium, when the stopcock is opened, the six atoms of gas are now distributed over a larger volume. To solve for the final partial pressures, use Boyle's law for each gas.

$$\text{For He: } P_2 = \frac{P_1V_1}{V_2} = P_{\text{He, initial}} \times \frac{X}{2X} = \frac{P_{\text{He, initial}}}{2}$$

The partial pressure of helium is exactly halved. The same result occurs with neon so that when the volume is doubled, the partial pressure is halved. Summarizing:

$$P_{\text{He, final}} = \frac{P_{\text{He, initial}}}{2}; \quad P_{\text{Ne, final}} = \frac{P_{\text{Ne, initial}}}{2}$$

**Gas Density, Molar Mass, and Reaction Stoichiometry**

50. Molar mass =  $\frac{dRT}{P}$  where d = density of gas in units of g/L.

$$\text{Molar mass} = \frac{3.164 \text{ g/L} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 273.2 \text{ K}}{1.000 \text{ atm}} = 70.93 \text{ g/mol}$$

The gas is diatomic, so the average atomic mass =  $70.93/2 = 35.47$  amu. From the periodic table, this is chlorine, and the identity of the gas is  $\text{Cl}_2$ .

51. If  $\text{Be}^{3+}$ , the formula is  $\text{Be}(\text{C}_5\text{H}_7\text{O}_2)_3$  and molar mass  $\approx 13.5 + 15(12) + 21(1) + 6(16) = 311$  g/mol. If  $\text{Be}^{2+}$ , the formula is  $\text{Be}(\text{C}_5\text{H}_7\text{O}_2)_2$  and molar mass  $\approx 9.0 + 10(12) + 14(1) + 4(16) = 207$  g/mol.

Data set I (molar mass =  $dRT/P$  and  $d = \text{mass}/V$ ):

$$\text{molar mass} = \frac{\text{mass} \times RT}{PV} = \frac{0.2022 \text{ g} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 286 \text{ K}}{(765.2 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}) \times (22.6 \times 10^{-3} \text{ L})} = 209 \text{ g/mol}$$

Data set II:

$$\text{molar mass} = \frac{\text{mass} \times RT}{PV} = \frac{0.2224 \text{ g} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 290. \text{ K}}{(764.6 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}) \times (26.0 \times 10^{-3} \text{ L})} = 202 \text{ g/mol}$$

These results are close to the expected value of 207 g/mol for  $\text{Be}(\text{C}_5\text{H}_7\text{O}_2)_2$ . Thus we conclude from these data that beryllium is a divalent element with an atomic weight (mass) of 9.0 g/mol.

52.  $d = P \times (\text{molar mass})/RT$ ; we need to determine the average molar mass of air. We get this by using the mole fraction information to determine the weighted value for the molar mass. If we have 1.000 mol of air:

$$\begin{aligned} \text{average molar mass} &= 0.78 \text{ mol N}_2 \times \frac{28.02 \text{ g N}_2}{\text{mol N}_2} + 0.21 \text{ mol O}_2 \times \frac{32.00 \text{ g O}_2}{\text{mol O}_2} \\ &\quad + 0.010 \text{ mol Ar} \times \frac{39.95 \text{ g Ar}}{\text{mol Ar}} = 28.98 = 29 \text{ g} \end{aligned}$$

$$d_{\text{air}} = \frac{1.00 \text{ atm} \times 29 \text{ g/mol}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 273 \text{ K}} = 1.3 \text{ g/L}$$

53. Rigid container: As temperature is increased, the gas molecules move with a faster average velocity. This results in more frequent and more forceful collisions, resulting in an increase in pressure. Density = mass/volume; the moles of gas are constant, and the volume of the container is constant, so density in this case must be temperature-independent (density is constant).

Flexible container: The flexible container is a constant-pressure container. Therefore, the final internal pressure will be unaffected by an increase in temperature. The density of the gas, however, will be affected because the container volume is affected. As T increases, there is an immediate increase in P inside the container. The container expands its volume to reduce the internal pressure back to the external pressure. We have the same mass of gas in a larger volume. Gas density will decrease in the flexible container as T increases.

54. We assume that 28.01 g/mol is the true value for the molar mass of N<sub>2</sub>. The value of 28.15 g/mol is the average molar mass of the amount of N<sub>2</sub> and Ar in air. Assume 100.00 mol of total gas present, and let x = the number of moles that are N<sub>2</sub> molecules. Then 100.00 - x = the number of moles that are Ar atoms. Solving:

$$28.15 = \frac{x(28.01) + (100.00 - x)(39.95)}{100.00}$$

$$2815 = (28.01)x + 3995 - (39.95)x, (11.94)x = 1180.$$

$$x = 98.83\% \text{ N}_2; \% \text{ Ar} = 100.00 - x = 1.17\% \text{ Ar}$$

$$\text{Ratio of moles of Ar to moles of N}_2 = \frac{1.17}{98.83} = 1.18 \times 10^{-2}.$$

55. Out of 100.0 g of compound, there are:

$$87.4 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 6.24 \text{ mol N}; \frac{6.24}{6.24} = 1.00$$

$$12.6 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 12.5 \text{ mol H}; \frac{12.5}{6.24} = 2.00$$

Empirical formula is NH<sub>2</sub>. P × (molar mass) = dRT, where d = density.

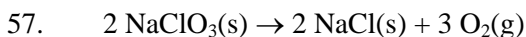
$$\text{Molar mass} = \frac{dRT}{P} = \frac{\frac{0.977 \text{ g}}{\text{L}} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 373 \text{ K}}{710. \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}} = 32.0 \text{ g/mol}$$

Empirical formula mass of NH<sub>2</sub> = 16.0 g. Therefore, the molecular formula is N<sub>2</sub>H<sub>4</sub>.

56. P × (molar mass) = dRT,  $d = \frac{\text{mass}}{\text{volume}}$ , P × (molar mass) =  $\frac{\text{mass}}{V} \times RT$

$$\text{Molar mass} = \frac{\text{mass} \times RT}{PV} = \frac{0.800 \text{ g} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 373 \text{ K}}{\left(750. \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}\right) \times 0.256 \text{ L}} = 96.9 \text{ g/mol}$$

$$\text{Mass of CHCl} \approx 12.0 + 1.0 + 35.5 = 48.5; \frac{96.9}{48.5} = 2.00; \text{ molecular formula is } \text{C}_2\text{H}_2\text{Cl}_2.$$



$$P_{\text{total}} = P_{\text{O}_2} + P_{\text{H}_2\text{O}}, P_{\text{O}_2} = P_{\text{total}} - P_{\text{H}_2\text{O}} = 734 \text{ torr} - 19.8 \text{ torr} = 714 \text{ torr}$$

$$n_{\text{O}_2} = \frac{P_{\text{O}_2} \times V}{RT} = \frac{\left(714 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}\right) \times 0.0572 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times (273 + 22) \text{ K}} = 2.22 \times 10^{-3} \text{ mol O}_2$$

$$\begin{aligned} \text{Mass NaClO}_3 \text{ decomposed} &= 2.22 \times 10^{-3} \text{ mol O}_2 \times \frac{2 \text{ mol NaClO}_3}{3 \text{ mol O}_2} \times \frac{106.44 \text{ g NaClO}_3}{\text{mol NaClO}_3} \\ &= 0.158 \text{ g NaClO}_3 \end{aligned}$$

$$\text{Mass \% NaClO}_3 = \frac{0.158 \text{ g}}{0.8765 \text{ g}} \times 100 = 18.0\%$$

58. For ammonia (in 1 minute):

$$n_{\text{NH}_3} = \frac{P_{\text{NH}_3} \times V_{\text{NH}_3}}{RT} = \frac{90. \text{ atm} \times 500. \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 496 \text{ K}} = 1.1 \times 10^3 \text{ mol NH}_3$$

$\text{NH}_3$  flows into the reactor at a rate of  $1.1 \times 10^3 \text{ mol/min}$ .

For  $\text{CO}_2$  (in 1 minute):

$$n_{\text{CO}_2} = \frac{P_{\text{CO}_2} \times V_{\text{CO}_2}}{RT} = \frac{45 \text{ atm} \times 600. \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 496 \text{ K}} = 6.6 \times 10^2 \text{ mol CO}_2$$

$\text{CO}_2$  flows into the reactor at  $6.6 \times 10^2 \text{ mol/min}$ .

To react completely with  $1.1 \times 10^3 \text{ mol NH}_3/\text{min}$ , we need:

$$\frac{1.1 \times 10^3 \text{ mol NH}_3}{\text{min}} \times \frac{1 \text{ mol CO}_2}{2 \text{ mol NH}_3} = 5.5 \times 10^2 \text{ mol CO}_2/\text{min}$$

Because 660 mol CO<sub>2</sub>/min are present, ammonia is the limiting reagent.

$$\frac{1.1 \times 10^3 \text{ mol NH}_3}{\text{min}} \times \frac{1 \text{ mol urea}}{2 \text{ mol NH}_3} \times \frac{60.06 \text{ g urea}}{\text{mol urea}} = 3.3 \times 10^4 \text{ g urea/min}$$

59.  $150 \text{ g (CH}_3)_2\text{N}_2\text{H}_2 \times \frac{1 \text{ mol (CH}_3)_2\text{N}_2\text{H}_2}{60.10 \text{ g}} \times \frac{3 \text{ mol N}_2}{\text{mol (CH}_3)_2\text{N}_2\text{H}_2} = 7.5 \text{ mol N}_2 \text{ produced}$

$$P_{\text{N}_2} = \frac{nRT}{V} = \frac{7.5 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 300. \text{ K}}{250 \text{ L}} = 0.74 \text{ atm}$$

We could do a similar calculation for P<sub>H<sub>2</sub>O</sub> and P<sub>CO<sub>2</sub></sub> and then calculate P<sub>total</sub> (= P<sub>N<sub>2</sub></sub> + P<sub>H<sub>2</sub>O</sub> + P<sub>CO<sub>2</sub></sub>). Or we can recognize that 9 total moles of gaseous products form for every mole of (CH<sub>3</sub>)<sub>2</sub>N<sub>2</sub>H<sub>2</sub> reacted. This is three times the moles of N<sub>2</sub> produced. Therefore, P<sub>total</sub> will be three times larger than P<sub>N<sub>2</sub></sub>. P<sub>total</sub> = 3 × P<sub>N<sub>2</sub></sub> = 3 × 0.74 atm = 2.2 atm.

60. Rigid container (constant volume): As reactants are converted to products, the moles of gas particles present decrease by one-half. As n decreases, the pressure will decrease (by one-half). Density is the mass per unit volume. Mass is conserved in a chemical reaction, so the density of the gas will not change because mass and volume do not change.

Flexible container (constant pressure): Pressure is constant because the container changes volume in order to keep a constant pressure. As the moles of gas particles decrease by a factor of 2, the volume of the container will decrease (by one-half). We have the same mass of gas in a smaller volume, so the gas density will increase (is doubled).

61. For NH<sub>3</sub>:  $P_2 = \frac{P_1 V_1}{V_2} = 0.500 \text{ atm} \times \frac{2.00 \text{ L}}{3.00 \text{ L}} = 0.333 \text{ atm}$

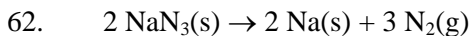
For O<sub>2</sub>:  $P_2 = \frac{P_1 V_1}{V_2} = 1.50 \text{ atm} \times \frac{1.00 \text{ L}}{3.00 \text{ L}} = 0.500 \text{ atm}$

After the stopcock is opened, V and T will be constant, so P ∝ n. The balanced equation requires:

$$\frac{n_{\text{O}_2}}{n_{\text{NH}_3}} = \frac{P_{\text{O}_2}}{P_{\text{NH}_3}} = \frac{5}{4} = 1.25$$

The actual ratio present is:  $\frac{P_{\text{O}_2}}{P_{\text{NH}_3}} = \frac{0.500 \text{ atm}}{0.333 \text{ atm}} = 1.50$

The actual ratio is larger than the required ratio, so  $\text{NH}_3$  in the denominator is limiting. Because equal moles of NO will be produced as  $\text{NH}_3$  reacted, the partial pressure of NO produced is 0.333 atm (the same as  $P_{\text{NH}_3}$  reacted).



$$n_{\text{N}_2} = \frac{PV}{RT} = \frac{1.00 \text{ atm} \times 70.0 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 273 \text{ K}} = 3.12 \text{ mol N}_2 \text{ are needed to fill the air bag.}$$

$$\text{Mass NaN}_3 \text{ reacted} = 3.12 \text{ mol N}_2 \times \frac{2 \text{ mol NaN}_3}{3 \text{ mol N}_2} \times \frac{65.02 \text{ g NaN}_3}{\text{mol NaN}_3} = 135 \text{ g NaN}_3$$

63. 
$$n_{\text{H}_2} = \frac{PV}{RT} = \frac{1.0 \text{ atm} \times \left[ 4800 \text{ m}^3 \times \left( \frac{100 \text{ cm}}{\text{m}} \right)^3 \times \frac{1 \text{ L}}{1000 \text{ cm}^3} \right]}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 273 \text{ K}} = 2.1 \times 10^5 \text{ mol}$$

$2.1 \times 10^5$  mol  $\text{H}_2$  are in the balloon. This is 80.% of the total amount of  $\text{H}_2$  that had to be generated:

$$0.80(\text{total mol H}_2) = 2.1 \times 10^5, \quad \text{total mol H}_2 = 2.6 \times 10^5 \text{ mol H}_2$$

$$2.6 \times 10^5 \text{ mol H}_2 \times \frac{1 \text{ mol Fe}}{\text{mol H}_2} \times \frac{55.85 \text{ g Fe}}{\text{mol Fe}} = 1.5 \times 10^7 \text{ g Fe}$$

$$2.6 \times 10^5 \text{ mol H}_2 \times \frac{1 \text{ mol H}_2\text{SO}_4}{\text{mol H}_2} \times \frac{98.09 \text{ g H}_2\text{SO}_4}{\text{mol H}_2\text{SO}_4} \times \frac{100 \text{ g reagent}}{98 \text{ g H}_2\text{SO}_4} = 2.6 \times 10^7 \text{ g of 98\% sulfuric acid}$$

64.  $0.2766 \text{ g CO}_2 \times \frac{12.011 \text{ g C}}{44.009 \text{ g CO}_2} = 7.549 \times 10^{-2} \text{ g C}; \quad \% \text{ C} = \frac{7.549 \times 10^{-2} \text{ g}}{0.1023 \text{ g}} \times 100 = 73.79\% \text{ C}$

$$0.0991 \text{ g H}_2\text{O} \times \frac{2.016 \text{ g H}}{18.02 \text{ g H}_2\text{O}} = 1.11 \times 10^{-2} \text{ g H}; \quad \% \text{ H} = \frac{1.11 \times 10^{-2} \text{ g}}{0.1023 \text{ g}} \times 100 = 10.9\% \text{ H}$$

$$PV = nRT, \quad n_{\text{N}_2} = \frac{PV}{RT} = \frac{1.00 \text{ atm} \times 27.6 \times 10^{-3} \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 273 \text{ K}} = 1.23 \times 10^{-3} \text{ mol N}_2$$

$$1.23 \times 10^{-3} \text{ mol N}_2 \times \frac{28.02 \text{ g N}_2}{\text{mol N}_2} = 3.45 \times 10^{-2} \text{ g nitrogen}$$

$$\text{Mass \% N} = \frac{3.45 \times 10^{-2} \text{ g}}{0.4831 \text{ g}} \times 100 = 7.14\% \text{ N}$$

$$\text{Mass \% O} = 100.00 - (73.79 + 10.9 + 7.14) = 8.2\% \text{ O}$$

Out of 100.00 g of compound, there are:

$$73.79 \text{ g C} \times \frac{1 \text{ mol}}{12.011 \text{ g}} = 6.144 \text{ mol C}; \quad 7.14 \text{ g N} \times \frac{1 \text{ mol}}{14.01 \text{ g}} = 0.510 \text{ mol N}$$

$$10.9 \text{ g H} \times \frac{1 \text{ mol}}{1.008 \text{ g}} = 10.8 \text{ mol H}; \quad 8.2 \text{ g O} \times \frac{1 \text{ mol}}{16.00 \text{ g}} = 0.51 \text{ mol O}$$

Dividing all values by 0.51 gives an empirical formula of  $\text{C}_{12}\text{H}_{21}\text{NO}$ .

$$\text{Molar mass} = \frac{dRT}{P} = \frac{\frac{4.02 \text{ g}}{\text{L}} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 400. \text{ K}}{256 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}} = 392 \text{ g/mol}$$

$$\text{Empirical formula mass of } \text{C}_{12}\text{H}_{21}\text{NO} \approx 195 \text{ g/mol and } \frac{392}{195} \approx 2.$$

Thus the molecular formula is  $\text{C}_{24}\text{H}_{42}\text{N}_2\text{O}_2$ .

65.  $2 \text{ NH}_3(\text{g}) \rightarrow \text{N}_2(\text{g}) + 3 \text{ H}_2(\text{g})$ ; as reactants are converted into products, we go from 2 moles of gaseous reactants to 4 moles of gaseous products (1 mol  $\text{N}_2$  + 3 mol  $\text{H}_2$ ). Because the moles of gas doubles as reactants are converted into products, the volume of the gases will double (at constant P and T).

$$PV = nRT, \quad P = \left( \frac{RT}{V} \right) n = (\text{constant})n; \quad \text{pressure is directly related to } n \text{ at constant } T \text{ and } V.$$

As the reaction occurs, the moles of gas will double, so the pressure will double. Because 1 mol of  $\text{N}_2$  is produced for every 2 mol of  $\text{NH}_3$  reacted,  $P_{\text{N}_2} = 1/2 P_{\text{NH}_3}^0$ . Owing to the 3 to 2 mole ratio in the balanced equation,  $P_{\text{H}_2} = 3/2 P_{\text{NH}_3}^0$ .

Note:  $P_{\text{total}} = P_{\text{H}_2} + P_{\text{N}_2} = 3/2 P_{\text{NH}_3}^0 + 1/2 P_{\text{NH}_3}^0 = 2 P_{\text{NH}_3}^0$ . As we said earlier, the total pressure doubles as reactants are completely converted into products for this reaction

66.  $1.00 \times 10^3 \text{ kg Mo} \times \frac{1000 \text{ g}}{\text{kg}} \times \frac{1 \text{ mol Mo}}{95.94 \text{ g Mo}} = 1.04 \times 10^4 \text{ mol Mo}$

$$1.04 \times 10^4 \text{ mol Mo} \times \frac{1 \text{ mol MoO}_3}{\text{mol Mo}} \times \frac{7/2 \text{ mol O}_2}{\text{mol MoO}_3} = 3.64 \times 10^4 \text{ mol O}_2$$

$$V_{\text{O}_2} = \frac{n_{\text{O}_2} RT}{P} = \frac{3.64 \times 10^4 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 290. \text{ K}}{1.00 \text{ atm}} = 8.66 \times 10^5 \text{ L of O}_2$$

$$8.66 \times 10^5 \text{ L O}_2 \times \frac{100 \text{ L air}}{21 \text{ L O}_2} = 4.1 \times 10^6 \text{ L air}$$

$$1.04 \times 10^4 \text{ mol Mo} \times \frac{3 \text{ mol H}_2}{\text{mol Mo}} = 3.12 \times 10^4 \text{ mol H}_2$$

$$V_{\text{H}_2} = \frac{3.12 \times 10^4 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 290. \text{ K}}{1.00 \text{ atm}} = 7.42 \times 10^5 \text{ L of H}_2$$

67.  $P_{\text{total}} = P_{\text{N}_2} + P_{\text{H}_2\text{O}}, P_{\text{N}_2} = 726 \text{ torr} - 23.8 \text{ torr} = 702 \text{ torr H} \frac{1 \text{ atm}}{760 \text{ torr}} = 0.924 \text{ atm}$

$$n_{\text{N}_2} = \frac{P_{\text{N}_2} \times V}{RT} = \frac{0.924 \text{ atm} \times 31.8 \times 10^{-3} \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298 \text{ K}} = 1.20 \times 10^{-3} \text{ mol N}_2$$

$$\text{Mass of N in compound} = 1.20 \times 10^{-3} \text{ mol N}_2 \times \frac{28.02 \text{ g N}_2}{\text{mol}} = 3.36 \times 10^{-2} \text{ g nitrogen}$$

$$\text{Mass \% N} = \frac{3.36 \times 10^{-2} \text{ g}}{0.253 \text{ g}} \times 100 = 13.3\% \text{ N}$$

68.  $10.10 \text{ atm} - 7.62 \text{ atm} = 2.48 \text{ atm}$  is the pressure of the amount of  $\text{F}_2$  reacted.

$$PV = nRT, V \text{ and } T \text{ are constant. } \frac{P}{n} = \text{constant, } \frac{P_1}{n_1} = \frac{P_2}{n_2} \text{ or } \frac{P_1}{P_2} = \frac{n_1}{n_2}$$

$$\frac{\text{Mol F}_2 \text{ reacted}}{\text{Mol Xe reacted}} = \frac{2.48 \text{ atm}}{1.24 \text{ atm}} = 2.00; \text{ so: } \text{Xe} + 2 \text{ F}_2 \rightarrow \text{XeF}_4$$

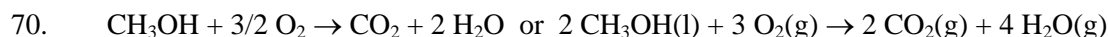
69. Because  $P$  and  $T$  are constant,  $V$  and  $n$  are directly proportional. The balanced equation requires 2 L of  $\text{H}_2$  to react with 1 L of  $\text{CO}$  (2 : 1 volume ratio due to 2 : 1 mole ratio in the balanced equation). The actual volume ratio present in 1 minute is  $16.0 \text{ L}/25.0 \text{ L} = 0.640$  (0.640 : 1). Because the actual volume ratio present is smaller than the required volume ratio,  $\text{H}_2$  is the limiting reactant. The volume of  $\text{CH}_3\text{OH}$  produced at STP will be one-half the volume of  $\text{H}_2$  reacted due to the 1 : 2 mole ratio in the balanced equation. In 1 minute,  $16.0 \text{ L}/2 = 8.00 \text{ L CH}_3\text{OH}$  are produced (theoretical yield).



$$n_{\text{CH}_3\text{OH}} = \frac{PV}{RT} = \frac{1.00 \text{ atm} \times 8.00 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 273 \text{ K}} = 0.357 \text{ mol CH}_3\text{OH in 1 minute}$$

$$0.357 \text{ mol CH}_3\text{OH} \times \frac{32.04 \text{ g CH}_3\text{OH}}{\text{mol CH}_3\text{OH}} = 11.4 \text{ g CH}_3\text{OH (theoretical yield per minute)}$$

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{5.30 \text{ g}}{11.4 \text{ g}} \times 100 = 46.5\% \text{ yield}$$



$$50.0 \text{ mL} \times \frac{0.850 \text{ g}}{\text{mL}} \times \frac{1 \text{ mol}}{32.04 \text{ g}} = 1.33 \text{ mol CH}_3\text{OH(l) available}$$

$$n_{\text{O}_2} = \frac{PV}{RT} = \frac{2.00 \text{ atm} \times 22.8 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 300. \text{ K}} = 1.85 \text{ mol O}_2 \text{ available}$$

$$1.33 \text{ mol CH}_3\text{OH} \times \frac{3 \text{ mol O}_2}{2 \text{ mol CH}_3\text{OH}} = 2.00 \text{ mol O}_2$$

2.00 mol O<sub>2</sub> are required to react completely with all the CH<sub>3</sub>OH available. We only have 1.85 mol O<sub>2</sub>, so O<sub>2</sub> is limiting.

$$1.85 \text{ mol O}_2 \times \frac{4 \text{ mol H}_2\text{O}}{3 \text{ mol O}_2} = 2.47 \text{ mol H}_2\text{O}$$

71. 
$$\text{Molar mass} = \frac{dRT}{P} = \frac{\frac{0.70902 \text{ g}}{\text{L}} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 273.2 \text{ K}}{1.000 \text{ atm}} = 15.90 \text{ g/mol}$$

15.90 g/mol is the average molar mass of the mixture of methane and helium. Assume 100.00 mol of total gas present, and let  $x$  = mol of CH<sub>4</sub> in the 100.00 mol mixture. This value of  $x$  is also equal to the volume percentage of CH<sub>4</sub> in 100.00 L of mixture because T and P are constant.

$$15.90 = \frac{x(16.04) + (100.00 - x)(4.003)}{100.00}, \quad 1590. = (16.04)x + 400.3 - (4.003)x$$

$$1190. = (12.04)x, \quad x = 98.84\% \text{ CH}_4 \text{ by volume; } \% \text{ He} = 100.00 - x = 1.16\% \text{ He by volume}$$

### Kinetic Molecular Theory and Real Gases

72. a. Containers ii, iv, vi, and viii have volumes twice those of containers i, iii, v, and vii. Containers iii, iv, vii, and viii have twice the number of molecules present than containers i, ii, v, and vi. The container with the lowest pressure will be the one which has the fewest moles of gas present in the largest volume (containers ii and vi both have the lowest P). The smallest container with the most moles of gas present will have the highest pressure (containers iii and vii both have the highest P). All the other containers (i, iv, v and viii) will have the same pressure between the two extremes. The order is  $ii = vi < i = iv = v = viii < iii = vii$ .
- b. All have the same average kinetic energy because the temperature is the same in each container. Only temperature determines the average kinetic energy.
- c. The least dense gas will be in container ii because it has the fewest of the lighter Ne atoms present in the largest volume. Container vii has the most dense gas because the largest number of the heavier Ar atoms are present in the smallest volume. To figure out the ordering for the other containers, we will calculate the relative density of each. In the table below,  $m_1$  equals the mass of Ne in container i,  $V_1$  equals the volume of container i, and  $d_1$  equals the density of the gas in container i.

Container	i	ii	iii	iv	v	vi	vii	viii
mass, volume	$m_1, V_1$	$m_1, 2V_1$	$2m_1, V_1$	$2m_1, 2V_1$	$2m_1, V_1$	$2m_1, 2V_1$	$4m_1, V_1$	$4m_1, 2V_1$
density $\left(\frac{\text{mass}}{\text{volume}}\right)$	$\frac{m_1}{V_1} = d_1$	$\frac{m_1}{2V_1} = \frac{1}{2}d_1$	$\frac{2m_1}{V_1} = 2d_1$	$\frac{2m_1}{2V_1} = d_1$	$\frac{2m_1}{V_1} = 2d_1$	$\frac{2m_1}{2V_1} = d_1$	$\frac{4m_1}{V_1} = 4d_1$	$\frac{4m_1}{2V_1} = 2d_1$

From the table, the order of gas density is  $ii < i = iv = vi < iii = v = viii < vii$ .

- d.  $\mu_{\text{rms}} = (3RT/M)^{1/2}$ ; the root mean square velocity only depends on the temperature and the molar mass. Because T is constant, the heavier argon molecules will have a slower root mean square velocity than the neon molecules. The order is  $v = vi = vii = viii < i = ii = iii = iv$ .
73. The kinetic molecular theory assumes that gas particles do not exert forces on each other and that gas particles are volumeless. Real gas particles do exert attractive forces for each other, and real gas particles do have volumes. A gas behaves most ideally at low pressures and high temperatures. The effect of attractive forces is minimized at high temperatures because the gas particles are moving very rapidly. At low pressure, the container volume is relatively large (P and V are inversely related), so the volume of the container taken up by the gas particles is negligible.
74. a. At constant temperature, the average kinetic energy of the He gas sample will equal the average kinetic energy of the  $\text{Cl}_2$  gas sample. In order for the average kinetic energies to be the same, the smaller He atoms must move at a faster average velocity than  $\text{Cl}_2$  molecules. Therefore, plot A, with the slower average velocity, would be for the  $\text{Cl}_2$

sample, and plot B would be for the He sample. Note the average velocity in each plot is a little past the top of the peak.

- b. As temperature increases, the average velocity of a gas will increase. Plot A would be for  $O_2(g)$  at 273 K, and plot B, with the faster average velocity, would be for  $O_2(g)$  at 1273 K.

Because a gas behaves more ideally at higher temperatures,  $O_2(g)$  at 1273 K would behave most ideally.

75. V, T, and P are all constant, so n must be constant. Because we have equal moles of gas in each container, gas B molecules must be heavier than gas A molecules.
- a. Both gas samples have the same number of molecules present (n is constant).
- b. Because T is constant,  $KE_{ave}$  must be the same for both gases ( $KE_{ave} = 3/2 RT$ ).
- c. The lighter gas A molecules will have the faster average velocity.
- d. The heavier gas B molecules do collide more forcefully, but gas A molecules, with the faster average velocity, collide more frequently. The end result is that P is constant between the two containers.
76. Boyle's law:  $P \propto 1/V$  at constant n and T

In the kinetic molecular theory (KMT), P is proportional to the collision frequency which is proportional to  $1/V$ . As the volume increases, there will be fewer collisions per unit area with the walls of the container, and pressure will decrease (Boyle's law).

Charles's law:  $V \propto T$  at constant n and P

When a gas is heated to a higher temperature, the velocities of the gas molecules increase and thus hit the walls of the container more often and with more force. In order to keep the pressure constant, the volume of the container must increase (this increases surface area, which decreases the number of collisions per unit area, which decreases the pressure). Therefore, volume and temperature are directly related at constant n and P (Charles's law).

Avogadro's law:  $V \propto n$  at constant P and T

As gas is added to a container (n increases), there will be an immediate increase in the number of gas particle collisions with the walls of the container. This results in an increase in pressure in the container. However, the container is such that it wants to keep the pressure constant. In order to keep pressure constant, the volume of the container increases in order to reduce the collision frequency, which reduces the pressure. V is directly related to n at constant P and T.

Dalton's law of partial pressure:  $P_{total} = P_1 + P_2 + P_3 + \dots$

The KMT assumes that gas particles are volumeless and that they exert no interparticle forces on each other. Gas molecules all behave the same way. Therefore, a mixture of gases behaves as one big gas sample. You can concentrate on the partial pressures of the individual components of the mixture, or you can collectively group all the gases together to determine the total pressure. One mole of an ideal gas behaves the same whether it is a pure gas or a mixture of gases.

P versus n relationship at constant V and T. This is a direct relationship. As gas is added to a container, there will be an increase in the collision frequency, resulting in an increase in pressure. P and n are directly related at constant V and T.

P versus T relationship at constant V and n. This is a direct relationship. As the temperature of the gas sample increases, the gas molecules move with a faster average velocity. This increases the gas collision frequency as well as increases the force of each gas particle collision. Both these result in an increase in pressure. Pressure and temperature are directly related at constant V and n.

77. Graham's law of effusion: 
$$\frac{\text{Rate}_1}{\text{Rate}_2} = \left( \frac{M_2}{M_1} \right)^{1/2}$$

Let Freon-12 = gas 1 and Freon-11 = gas 2:

$$\frac{1.07}{1.00} = \left( \frac{137.4}{M_1} \right)^{1/2}, \quad 1.14 = \frac{137.4}{M_1}, \quad M_1 = 121 \text{ g/mol}$$

The molar mass of  $\text{CF}_2\text{Cl}_2$  is equal to 121 g/mol, so Freon-12 is  $\text{CF}_2\text{Cl}_2$ .

78. 
$$\frac{\text{Rate}_1}{\text{Rate}_2} = \left( \frac{M_2}{M_1} \right)^{1/2}, \quad \frac{{}^{12}\text{C}^{17}\text{O}}{{}^{12}\text{C}^{18}\text{O}} = \left( \frac{30.0}{29.0} \right)^{1/2} = 1.02; \quad \frac{{}^{12}\text{C}^{16}\text{O}}{{}^{12}\text{C}^{18}\text{O}} = \left( \frac{30.0}{28.0} \right)^{1/2} = 1.04$$

The relative rates of effusion of  ${}^{12}\text{C}^{16}\text{O} : {}^{12}\text{C}^{17}\text{O} : {}^{12}\text{C}^{18}\text{O}$  are 1.04 : 1.02 : 1.00.

Advantage:  $\text{CO}_2$  isn't as toxic as CO.

Major disadvantages of using  $\text{CO}_2$  instead of CO:

1. Can get a mixture of oxygen isotopes in  $\text{CO}_2$ .
2. Some species, for example,  ${}^{12}\text{C}^{16}\text{O}^{18}\text{O}$  and  ${}^{12}\text{C}^{17}\text{O}_2$ , would effuse (gaseously diffuse) at about the same rate because the masses are about equal. Thus some species cannot be separated from each other.

79. The number of gas particles is constant, so at constant moles of gas, either a temperature change or a pressure change results in the smaller volume. If the temperature is constant, an increase in the external pressure would cause the volume to decrease. Gases are mostly empty space so gases are easily compressible.

If the pressure is constant, a decrease in temperature would cause the volume to decrease. As the temperature is lowered, the gas particles move with a slower average velocity and don't collide with the container walls as frequently and as forcefully. As a result, the internal pressure decreases. In order to keep the pressure constant, the volume of the container must decrease in order to increase the gas particle collisions per unit area.

80. In this situation, the volume has increased by a factor of two. One way to double the volume of a container at constant pressure and temperature is to double the number of moles of gas particles present. As gas particles are added, more collisions per unit area occur and the internal pressure increases. In order to keep the pressure constant, the container volume must increase.

Another way to double the volume of a container at constant pressure and moles of gas is to double the absolute temperature. As temperature increases, the gas molecules collide more frequently with the walls of the container. In order to keep pressure constant, the container volume must increase.

The last variable which can be changed is pressure. If the external pressure exerted on the container is halved, the volume will double (assuming constant temperature and moles). As the external pressure applied is reduced, the volume of the container must increase in order to equalize the higher internal pressure with the lower external applied pressure.

$$81. \frac{\text{Rate}_1}{\text{Rate}_2} = \left( \frac{M_2}{M_1} \right)^{1/2}; \quad \text{rate}_1 = \frac{24.0 \text{ mL}}{\text{min}}; \quad \text{rate}_2 = \frac{47.8 \text{ mL}}{\text{min}}; \quad M_2 = \frac{16.04 \text{ g}}{\text{mol}}; \quad M_1 = ?$$

$$\frac{24.0}{47.8} = \left( \frac{16.04}{M_1} \right)^{1/2} = 0.502, \quad 16.04 = (0.502)^2 \times M_1, \quad M_1 = \frac{16.04}{0.252} = \frac{63.7 \text{ g}}{\text{mol}}$$

$$82. \frac{\text{Rate}_1}{\text{Rate}_2} = \left( \frac{M_2}{M_1} \right)^{1/2} \quad \text{where } M = \text{molar mass}; \quad \text{let gas (1) = He, gas (2) = Cl}_2.$$

$$\frac{\frac{1.0 \text{ L}}{4.5 \text{ min}}}{\frac{1.0 \text{ L}}{t}} = \left( \frac{70.90}{4.003} \right)^{1/2}, \quad \frac{t}{4.5 \text{ min}} = 4.209, \quad t = 19 \text{ min}$$

83. a. They will all have the same average kinetic energy because they are all at the same temperature. Average kinetic energy depends only on temperature.
- b. Flask C; at constant T,  $u_{\text{rms}} \propto (1/M)^{1/2}$ . In general, the lighter the gas molecules, the greater is the root mean square velocity (at constant T).
- c. Flask A: collision frequency is proportional to average velocity  $\times n/V$  (as the average velocity doubles, the number of collisions will double, and as the number of molecules in the container doubles, the number of collisions again doubles). At constant T and V, n is

proportional to  $P$ , and average velocity is proportional to  $(1/M)^{1/2}$ . We use these relationships and the data in the exercise to determine the following relative values.

	n (relative)	$u_{\text{avg}}$ (relative)	Coll. Freq. (relative) = $n \times u_{\text{avg}}$
A	1.0	1.0	1.0
B	0.33	1.0	0.33
C	0.13	3.7	0.48

84.	a	b	c	d
Avg. KE ( $KE_{\text{avg}} \propto T$ )	inc	dec	same	same
$u_{\text{rms}}$ ( $u_{\text{rms}}^2 \propto T$ )	inc	dec	same	same
Coll. freq. gas	inc	dec	inc	inc
Coll. freq. wall	inc	dec	inc	inc
Impact E (impact E $\propto KE \propto T$ )	inc	dec	same	same

Both collision frequencies are proportional to the root mean square velocity (as velocity increases, it takes less time to move to the next collision) and the quantity  $n/V$  (as molecules per volume increases, collision frequency increases).

85. No; at each temperature there is a distribution of energies. Similarly, there is a distribution of velocities at any specific temperature (see Figs. 5.15 to 5.17 of the text). Note that the major reason there is a distribution of kinetic energies is because there is a distribution of velocities for any gas sample at some temperature.

86. a. All the gases have the same average kinetic energy because they are all at the same temperature [ $KE_{\text{ave}} = (3/2)RT$ ].

b. At constant  $T$ , the lighter the gas molecule, the faster is the average velocity [ $u_{\text{ave}} \propto (1/M)^{1/2}$ ].

Xe (131.3 g/mol) < Cl<sub>2</sub> (70.90 g/mol) < O<sub>2</sub> (32.00 g/mol) < H<sub>2</sub> (2.016 g/mol)  
 slowest fastest

c. At constant  $T$ , the lighter H<sub>2</sub> molecules have a faster average velocity than the heavier O<sub>2</sub> molecules. As temperature increases, the average velocity of the gas molecules increases. Separate samples of H<sub>2</sub> and O<sub>2</sub> can only have the same average velocities if the temperature of the O<sub>2</sub> sample is greater than the temperature of the H<sub>2</sub> sample.

87. a.  $PV = nRT$

$$P = \frac{nRT}{V} = \frac{0.5000 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times (25.0 + 273.2) \text{ K}}{1.0000 \text{ L}} = 12.24 \text{ atm}$$

b.  $\left[ P + a \left( \frac{n}{V} \right)^2 \right] (V - nb) = nRT$ ; for  $N_2$ :  $a = 1.39 \text{ atm L}^2/\text{mol}^2$  and  $b = 0.0391 \text{ L/mol}$

$$\left[ P + 1.39 \left( \frac{0.5000}{1.0000} \right)^2 \text{ atm} \right] (1.0000 \text{ L} - 0.5000 \times 0.0391 \text{ L}) = 12.24 \text{ L atm}$$

$$(P + 0.348 \text{ atm})(0.9805 \text{ L}) = 12.24 \text{ L atm}$$

$$P = \frac{12.24 \text{ L atm}}{0.9805 \text{ L}} - 0.348 \text{ atm} = 12.48 - 0.348 = 12.13 \text{ atm}$$

c. The ideal gas law is high by 0.11 atm, or  $\frac{0.11}{12.13} \times 100 = 0.91\%$ .

88. a.  $P = \frac{nRT}{V} = \frac{0.5000 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298.2 \text{ K}}{10.000 \text{ L}} = 1.224 \text{ atm}$

b.  $\left[ P + a \left( \frac{n}{V} \right)^2 \right] (V - nb) = nRT$ ; for  $N_2$ :  $a = 1.39 \text{ atm L}^2/\text{mol}^2$  and  $b = 0.0391 \text{ L/mol}$

$$\left[ P + 1.39 \left( \frac{0.5000}{10.000} \right)^2 \text{ atm} \right] (10.000 \text{ L} - 0.5000 \times 0.0391 \text{ L}) = 12.24 \text{ L atm}$$

$$(P + 0.00348 \text{ atm})(10.000 \text{ L} - 0.0196 \text{ L}) = 12.24 \text{ L atm}$$

$$P + 0.00348 \text{ atm} = \frac{12.24 \text{ L atm}}{9.980 \text{ L}} = 1.226 \text{ atm}, P = 1.226 - 0.00348 = 1.223 \text{ atm}$$

c. The results agree to  $\pm 0.001 \text{ atm}$  (0.08%).

d. In Exercise 87 the pressure is relatively high and there is significant disagreement. In Exercise 88 the pressure is around 1 atm and both gas laws show better agreement. The ideal gas law is valid at relatively low pressures.

89.  $(KE)_{\text{avg}} = 3/2 RT$ ; KE depends only on temperature. At each temperature  $CH_4$  and  $N_2$  will have the same average KE. For energy units of joules (J), use  $R = 8.3145 \text{ J K}^{-1} \text{ mol}^{-1}$ . To determine average KE per molecule, divide by Avogadro's number,  $6.022 \times 10^{23}$  molecules/mol.

$$\text{At } 273 \text{ K: } (KE)_{\text{avg}} = \frac{3}{2} \times \frac{8.3145 \text{ J}}{\text{K mol}} \times 273 \text{ K} = 3.40 \times 10^3 \text{ J/mol} = 5.65 \times 10^{-21} \text{ J/molecule}$$

$$\text{At } 546 \text{ K: } (\text{KE})_{\text{avg}} = \frac{3}{2} \times \frac{8.3145 \text{ J}}{\text{K mol}} \times 546 \text{ K} = 6.81 \times 10^3 \text{ J/mol} = 1.13 \times 10^{-20} \text{ J/molecule}$$

$$90. \quad u_{\text{rms}} = \left( \frac{3RT}{M} \right)^{1/2}, \text{ where } R = \frac{8.3145 \text{ J}}{\text{K mol}} \text{ and } M = \text{molar mass in kg}$$

For  $\text{CH}_4$ :  $M = 1.604 \times 10^{-2} \text{ kg}$  and for  $\text{N}_2$ ,  $M = 2.802 \times 10^{-2} \text{ kg}$ .

$$\text{For } \text{CH}_4 \text{ at } 273 \text{ K: } u_{\text{rms}} = \left( \frac{3 \times \frac{8.3145 \text{ J}}{\text{K mol}} \times 273 \text{ K}}{1.604 \times 10^{-2} \text{ kg/mol}} \right)^{1/2} = 652 \text{ m/s}$$

At 546 K:  $u_{\text{rms}}$  for  $\text{CH}_4$  is 921 m/s.

For  $\text{N}_2$ :  $u_{\text{rms}} = 493 \text{ m/s}$  at 273 K and 697 m/s at 546 K.

$$91. \quad u_{\text{rms}} = \left( \frac{3RT}{M} \right)^{1/2} = \left[ \frac{3 \left( \frac{8.3145 \text{ kg m}^2}{\text{s}^2 \text{ K mol}} \right) (227 + 273) \text{ K}}{28.02 \times 10^{-3} \text{ kg/mol}} \right]^{1/2} = 667 \text{ m/s}$$

$$u_{\text{mp}} = \left( \frac{2RT}{M} \right)^{1/2} = \left[ \frac{2 \left( \frac{8.3145 \text{ kg m}^2}{\text{s}^2 \text{ K mol}} \right) (500. \text{ K})}{28.02 \times 10^{-3} \text{ kg/mol}} \right]^{1/2} = 545 \text{ m/s}$$

$$u_{\text{avg}} = \left( \frac{8RT}{\pi M} \right)^{1/2} = \left[ \frac{8 \left( \frac{8.3145 \text{ kg m}^2}{\text{s}^2 \text{ K mol}} \right) (500. \text{ K})}{\pi (28.02 \times 10^{-3} \text{ kg/mol})} \right]^{1/2} = 615 \text{ m/s}$$

$$92. \quad \text{KE}_{\text{ave}} = 3/2 \text{ RT per mol; } \text{KE}_{\text{ave}} = 3/2 \text{ k}_B \text{T per molecule}$$

$$\text{KE}_{\text{total}} = 3/2 \times (1.3807 \times 10^{-23} \text{ J/K}) \times 300. \text{ K} \times (1.00 \times 10^{20} \text{ molecules}) = 0.621 \text{ J}$$

$$93. \quad \text{The values of } a \text{ are: } \text{H}_2, \frac{0.244 \text{ atm L}^2}{\text{mol}^2}; \text{CO}_2, 3.59; \text{N}_2, 1.39; \text{CH}_4, 2.25$$

Because  $a$  is a measure of intermolecular attractions, the attractions are greatest for  $\text{CO}_2$ .



94. The van der Waals constant  $b$  is a measure of the size of the molecule. Thus  $C_3H_8$  should have the largest value of  $b$  because it has the largest molar mass (size).
95. The pressure measured for real gases is too low compared to ideal gases. This is due to the attractions gas particles do have for each other; these attractions “hold” them back from hitting the container walls as forcefully. To make up for this slight decrease in pressure for real gases, a factor is added to the measured pressure. The measured volume is too large. A fraction of the space of the container volume is taken up by the volume of the molecules themselves. Therefore, the actual volume available to real gas molecules is slightly less than the container volume. A term is subtracted from the container volume to correct for the volume taken up by real gas molecules.

$$96. \quad f(u) = 4\pi \left( \frac{m}{2\pi k_B T} \right)^{3/2} u^2 e^{(-mu^2/2k_B T)}$$

As  $u \rightarrow 0$ ,  $e^{(-mu^2/2k_B T)} \rightarrow e^0 = 1$ ; at small values of  $u$ , the  $u^2$  term causes the function to increase. At large values of  $u$ , the exponent term,  $-mu^2/2k_B T$ , is a large negative number, and  $e$  raised to a large negative number causes the function to decrease. As  $u \rightarrow \infty$ ,  $e^{-\infty} \rightarrow 0$ .

97. Intermolecular collision frequency =  $Z = 4 \frac{N}{V} d^2 \left( \frac{\pi RT}{M} \right)^{1/2}$ , where  $d$  = diameter of He atom

$$\frac{n}{V} = \frac{P}{RT} = \frac{3.0 \text{ atm}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 300. \text{ K}} = 0.12 \text{ mol/L}$$

$$\frac{N}{V} = \frac{0.12 \text{ mol}}{\text{L}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} \times \frac{1000 \text{ L}}{\text{m}^3} = \frac{7.2 \times 10^{25} \text{ molecules}}{\text{m}^3}$$

$$Z = 4 \times \frac{7.2 \times 10^{25} \text{ molecules}}{\text{m}^3} \times (50. \times 10^{-12} \text{ m})^2 \times \left( \frac{\pi(8.3145)(300.)}{4.00 \times 10^{-3}} \right)^{1/2} \\ = 1.0 \times 10^9 \text{ collisions/s}$$

$$\text{Mean free path} = \lambda = \frac{u_{\text{avg}}}{Z}; \quad u_{\text{avg}} = \left( \frac{8 RT}{\pi M} \right)^{1/2} = 1260 \text{ m/s}; \quad \lambda = \frac{1260 \text{ m/s}}{1.0 \times 10^9 \text{ s}^{-1}} = 1.3 \times 10^{-6} \text{ m}$$

98.  $\Delta(mu) = 2mu$  and  $u \propto (T/M)^{1/2}$ ;  $\frac{\Delta(mu)_{77}}{\Delta(mu)_{27}} = \frac{2m \left( \frac{350. \text{ K}}{M} \right)^{1/2}}{2m \left( \frac{300. \text{ K}}{M} \right)^{1/2}} = \left( \frac{350.}{300.} \right)^{1/2} = 1.08$

The change in momentum is 1.08 times greater for Ar at 77°C than for Ar at 27°C.

$$Z_A = A \frac{N}{V} \left( \frac{RT}{2\pi M} \right)^{1/2}; \quad \frac{Z_{77}}{Z_{27}} = \left( \frac{T_{77}}{T_{27}} \right)^{1/2}, \quad \frac{Z_{77}}{Z_{27}} = \left( \frac{350.}{300.} \right)^{1/2} = 1.08$$

There are 1.08 times as many impacts per second for Ar at 77°C as for Ar at 27°C.

99. The force per impact is proportional to  $\Delta(\mu) = 2mu$ . Because  $m \propto M$ , the molar mass, and  $u \propto (1/M)^{1/2}$  at constant  $T$ , the force per impact at constant  $T$  is proportional to  $M \times (1/M)^{1/2} = \sqrt{M}$ .

$$\frac{\text{Impactforce}(\text{H}_2)}{\text{Impactforce}(\text{He})} = \sqrt{\frac{M_{\text{H}_2}}{M_{\text{He}}}} = \sqrt{\frac{2.016}{4.003}} = 0.7097$$

100.  $\frac{\text{Diffusion rate } ^{235}\text{UF}_6}{\text{Diffusion rate } ^{238}\text{UF}_6} = 1.0043$  (See Section 5.7 of the text.)

$$\frac{^{235}\text{UF}_6}{^{238}\text{UF}_6} \times (1.0043)^{100} = \frac{1526}{1.000 \times 10^5 - 1526}, \quad \frac{^{235}\text{UF}_6}{^{238}\text{UF}_6} \times 1.5358 = \frac{1526}{98500}$$

$$\frac{^{235}\text{UF}_6}{^{238}\text{UF}_6} = 1.01 \times 10^{-2} = \text{initial } ^{235}\text{U to } ^{238}\text{U atom ratio}$$

101.  $\Delta(\mu) = 2mu = \text{change in momentum per impact}$ . Because  $m$  is proportional to  $M$ , the molar mass, and  $u$  is proportional to  $(T/M)^{1/2}$ :

$$\Delta(\mu)_{\text{O}_2} \propto 2M_{\text{O}_2} \left( \frac{T}{M_{\text{O}_2}} \right)^{1/2} \quad \text{and} \quad \Delta(\mu)_{\text{He}} \propto 2M_{\text{He}} \left( \frac{T}{M_{\text{He}}} \right)^{1/2}$$

$$\frac{\Delta(\mu)_{\text{O}_2}}{\Delta(\mu)_{\text{He}}} = \frac{2M_{\text{O}_2} \left( \frac{T}{M_{\text{O}_2}} \right)^{1/2}}{2M_{\text{He}} \left( \frac{T}{M_{\text{He}}} \right)^{1/2}} = \frac{M_{\text{O}_2}}{M_{\text{He}}} \left( \frac{M_{\text{He}}}{M_{\text{O}_2}} \right)^{1/2} = \frac{31.998}{4.003} \left( \frac{4.003}{31.998} \right)^{1/2} = 2.827$$

The change in momentum per impact is 2.827 times larger for  $\text{O}_2$  molecules than for He atoms.

$$Z_A = A \frac{N}{V} \left( \frac{RT}{2\pi M} \right)^{1/2} = \text{collision rate}$$

$$\frac{Z_{\text{O}_2}}{Z_{\text{He}}} = \frac{A \left( \frac{N}{V} \right) \left( \frac{RT}{2\pi M_{\text{O}_2}} \right)^{1/2}}{A \left( \frac{N}{V} \right) \left( \frac{RT}{2\pi M_{\text{He}}} \right)^{1/2}} = \frac{\left( \frac{1}{M_{\text{O}_2}} \right)^{1/2}}{\left( \frac{1}{M_{\text{He}}} \right)^{1/2}} = 0.3537, \quad \frac{Z_{\text{He}}}{Z_{\text{O}_2}} = 2.827$$

There are 2.827 times as many impacts per second for He as for  $\text{O}_2$ .

**Atmospheric Chemistry**

102.  $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{NO}(\text{g})$ , automobile combustion or formed by lightning

$2 \text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{NO}_2(\text{g})$ , reaction with atmospheric  $\text{O}_2$

$2 \text{NO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{HNO}_3(\text{aq}) + \text{HNO}_2(\text{aq})$ , reaction with atmospheric  $\text{H}_2\text{O}$

$\text{S}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g})$ , combustion of coal

$2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{SO}_3(\text{g})$ , reaction with atmospheric  $\text{O}_2$

$\text{H}_2\text{O}(\text{l}) + \text{SO}_3(\text{g}) \rightarrow \text{H}_2\text{SO}_4(\text{aq})$ , reaction with atmospheric  $\text{H}_2\text{O}$

$2 \text{HNO}_3(\text{aq}) + \text{CaCO}_3(\text{s}) \rightarrow \text{Ca}(\text{NO}_3)_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$

$\text{H}_2\text{SO}_4(\text{aq}) + \text{CaCO}_3(\text{s}) \rightarrow \text{CaSO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$

103. a. If we have  $1.0 \times 10^6$  L of air, then there are  $3.0 \times 10^2$  L of CO.

$$P_{\text{CO}} = \chi_{\text{CO}} P_{\text{total}}; \chi_{\text{CO}} = \frac{V_{\text{CO}}}{V_{\text{total}}} \text{ because } V \% n; P_{\text{CO}} = \frac{3.0 \times 10^2}{1.0 \times 10^6} \times 628 \text{ torr} = 0.19 \text{ torr}$$

b.  $n_{\text{CO}} = \frac{P_{\text{CO}} V}{RT}$ ; Assuming  $1.0 \text{ m}^3$  air,  $1 \text{ m}^3 = 1000 \text{ L}$ :

$$n_{\text{CO}} = \frac{0.19 \text{ atm} \times (1.0 \times 10^3 \text{ L})}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 273 \text{ K}} = 1.1 \times 10^{-2} \text{ mol CO}$$

$$1.1 \times 10^{-2} \text{ mol} \times \frac{6.02 \times 10^{23} \text{ molecules}}{\text{mol}} = 6.6 \times 10^{21} \text{ CO molecules in } 1.0 \text{ m}^3 \text{ of air}$$

c.  $\frac{6.6 \times 10^{21} \text{ molecules}}{\text{m}^3} \times \left( \frac{1 \text{ m}}{100 \text{ cm}} \right)^3 = \frac{6.6 \times 10^{15} \text{ molecules CO}}{\text{cm}^3}$

104.  $\chi_{\text{He}} = 5.24 \times 10^{-6}$  from Table 5.4.  $P_{\text{He}} = \chi_{\text{He}} \times P_{\text{total}} = 5.24 \times 10^{-6} \times 1.0 \text{ atm} = 5.2 \times 10^{-6} \text{ atm}$

$$\frac{n}{V} = \frac{P}{RT} = \frac{5.2 \times 10^{-6} \text{ atm}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298 \text{ K}} = 2.1 \times 10^{-7} \text{ mol He/L}$$

$$\frac{2.1 \times 10^{-7} \text{ mol}}{\text{L}} \times \frac{1 \text{ L}}{1000 \text{ cm}^3} \times \frac{6.022 \times 10^{23} \text{ atoms}}{\text{mol}} = 1.3 \times 10^{14} \text{ atoms He/cm}^3$$

105. For benzene:

$$89.6 \times 10^{-9} \text{ g} \times \frac{1 \text{ mol}}{78.11 \text{ g}} = 1.15 \times 10^{-9} \text{ mol benzene}$$

$$V_{\text{benzene}} = \frac{n_{\text{benzene}} RT}{P} = \frac{1.15 \times 10^{-9} \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 296 \text{ K}}{748 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}} = 2.84 \times 10^{-8} \text{ L}$$

$$\text{Mixing ratio} = \frac{2.84 \times 10^{-8} \text{ L}}{3.00 \text{ L}} \times 10^6 = 9.47 \times 10^{-3} \text{ ppmv}$$

$$\text{or ppbv} = \frac{\text{vol of X} \times 10^9}{\text{total vol}} = \frac{2.84 \times 10^{-8} \text{ L}}{3.00 \text{ L}} \times 10^9 = 9.47 \text{ ppbv}$$

$$\begin{aligned} \frac{1.15 \times 10^{-9} \text{ mol benzene}}{3.00 \text{ L}} \times \frac{1 \text{ L}}{1000 \text{ cm}^3} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} \\ = 2.31 \times 10^{11} \text{ molecules benzene/cm}^3 \end{aligned}$$

For toluene:

$$153 \times 10^{-9} \text{ g C}_7\text{H}_8 \times \frac{1 \text{ mol}}{92.13 \text{ g}} = 1.66 \times 10^{-9} \text{ mol toluene}$$

$$V_{\text{toluene}} = \frac{n_{\text{toluene}} RT}{P} = \frac{1.66 \times 10^{-9} \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 296 \text{ K}}{748 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}} = 4.10 \times 10^{-8} \text{ L}$$

$$\text{Mixing ratio} = \frac{4.10 \times 10^{-8} \text{ L}}{3.00 \text{ L}} \times 10^6 = 1.37 \times 10^{-2} \text{ ppmv (or 13.7 ppbv)}$$

$$\begin{aligned} \frac{1.66 \times 10^{-9} \text{ mol toluene}}{3.00 \text{ L}} \times \frac{1 \text{ L}}{1000 \text{ cm}^3} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} \\ = 3.33 \times 10^{11} \text{ molecules toluene/cm}^3 \end{aligned}$$

### Additional Exercises

106. Statements a, c, and e are true. For statement b, if temperature is constant, then the average kinetic energy will be constant no matter what the identity of the gas ( $KE_{\text{ave}} = 3/2 RT$ ). For statement d, as T increases, the average velocity of the gas particles increases. When gas particles are moving faster, the effect of interparticle interactions is minimized. For statement f, the KMT predicts that P is directly related to T at constant V and n. As T increases, the gas molecules move faster, on average, resulting in more frequent and more forceful collisions. This leads to an increase in P.

107. At constant T and P, Avogadro's law applies; that is, equal volumes contain equal moles of molecules. In terms of balanced equations, we can say that mole ratios and volume ratios between the various reactants and products will be equal to each other.  $\text{Br}_2 + 3 \text{F}_2 \rightarrow 2 \text{X}$ ; 2 moles of X must contain 2 moles of Br and 6 moles of F; X must have the formula  $\text{BrF}_3$  for a balanced equation.

108. The partial pressure of  $\text{CO}_2$  that reacted is  $740. - 390. = 350.$  torr. Thus the number of moles of  $\text{CO}_2$  that reacts is given by:

$$n = \frac{PV}{RT} = \frac{\frac{350.}{760} \text{ atm} \times 3.00 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 293 \text{ K}} = 5.75 \times 10^{-2} \text{ mol CO}_2$$

$$5.75 \times 10^{-2} \text{ mol CO}_2 \times \frac{1 \text{ mol MgO}}{1 \text{ mol CO}_2} \times \frac{40.31 \text{ g MgO}}{\text{mol MgO}} = 2.32 \text{ g MgO}$$

$$\text{Mass \% MgO} = \frac{2.32 \text{ g}}{2.85 \text{ g}} \times 100 = 81.4\% \text{ MgO}$$

109. Processes a, c, and d will all result in a doubling of the pressure. Process a has the effect of halving the volume, which would double the pressure (Boyle's law). Process c doubles the pressure because the absolute temperature is doubled (from 200. K to 400. K). Process d doubles the pressure because the moles of gas are doubled (28 g  $\text{N}_2$  is 1 mol of  $\text{N}_2$ ). Process b won't double the pressure since the absolute temperature is not doubled (303 K to 333 K).

$$110. \quad 14.1 \times 10^2 \text{ in Hg} \cdot \text{in}^3 \times \frac{2.54 \text{ cm}}{\text{in}} \times \frac{10 \text{ mm}}{1 \text{ cm}} \times \frac{1 \text{ atm}}{760 \text{ mm}} \times \left( \frac{2.54 \text{ cm}}{\text{in}} \right)^3 \times \frac{1 \text{ L}}{1000 \text{ cm}^3} \\ = 0.772 \text{ atm L}$$

Boyle's law:  $PV = k$ , where  $k = nRT$ ; from Example 5.1 of the text, the k values are around 22 atm L. Because  $k = nRT$ , we can assume that Boyle's data and the Example 5.1 data were taken at different temperatures and/or had different sample sizes (different moles).

$$111. \quad PV = nRT, V \text{ and } T \text{ are constant. } \frac{P_1}{n_1} = \frac{P_2}{n_2}, \frac{P_2}{P_1} = \frac{n_2}{n_1}$$

We will do this limiting-reagent problem using an alternative method than described in Chapter 3. Let's calculate the partial pressure of  $C_3H_3N$  that can be produced from each of the starting materials assuming each reactant is limiting. The reactant that produces the smallest amount of product will run out first and is the limiting reagent.

$$P_{C_3H_3N} = 0.500 \text{ MPa} \times \frac{2 \text{ MPa } C_3H_3N}{2 \text{ MPa } C_3H_6} = 0.500 \text{ MPa if } C_3H_6 \text{ is limiting}$$

$$P_{C_3H_3N} = 0.800 \text{ MPa} \times \frac{2 \text{ MPa } C_3H_3N}{2 \text{ MPa } NH_3} = 0.800 \text{ MPa if } NH_3 \text{ is limiting}$$

$$P_{C_3H_3N} = 1.500 \text{ MPa} \times \frac{2 \text{ MPa } C_3H_3N}{3 \text{ MPa } O_2} = 1.000 \text{ MPa if } O_2 \text{ is limiting}$$

$C_3H_6$  is limiting. Although more product could be produced from  $NH_3$  and  $O_2$ , there is only enough  $C_3H_6$  to produce 0.500 MPa of  $C_3H_3N$ . The partial pressure of  $C_3H_3N$  in atmospheres after the reaction is:

$$0.500 \times 10^6 \text{ Pa} \times \frac{1 \text{ atm}}{1.013 \times 10^5 \text{ Pa}} = 4.94 \text{ atm}$$

$$n = \frac{PV}{RT} = \frac{4.94 \text{ atm} \times 150. \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298 \text{ K}} = 30.3 \text{ mol } C_3H_3N$$

$$30.3 \text{ mol} \times \frac{53.06 \text{ g}}{\text{mol}} = 1.61 \times 10^3 \text{ g } C_3H_3N \text{ can be produced.}$$

$$112. \quad 750. \text{ mL juice} \times \frac{12 \text{ mL } C_2H_5OH}{100 \text{ mL juice}} = 90. \text{ mL } C_2H_5OH \text{ present}$$

$$90. \text{ mL } C_2H_5OH \times \frac{0.79 \text{ g } C_2H_5OH}{\text{mL } C_2H_5OH} \times \frac{1 \text{ mol } C_2H_5OH}{46.07 \text{ g } C_2H_5OH} \times \frac{2 \text{ mol } CO_2}{2 \text{ mol } C_2H_5OH} = 1.5 \text{ mol } CO_2$$

The  $CO_2$  will occupy  $(825 - 750. =) 75 \text{ mL}$  not occupied by the liquid (headspace).

$$P_{CO_2} = \frac{n_{CO_2} RT}{V} = \frac{1.5 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298 \text{ K}}{75 \times 10^{-3} \text{ L}} = 490 \text{ atm}$$

Actually, enough  $CO_2$  will dissolve in the wine to lower the pressure of  $CO_2$  to a much more reasonable value.

$$113. \quad 0.050 \text{ mL} \times \frac{1.149 \text{ g}}{\text{mL}} \times \frac{1 \text{ mol O}_2}{32.0 \text{ g}} = 1.8 \times 10^{-3} \text{ mol O}_2$$

$$V = \frac{nRT}{P} = \frac{1.8 \times 10^{-3} \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 310. \text{ K}}{1.0 \text{ atm}} = 4.6 \times 10^{-2} \text{ L} = 46 \text{ mL}$$

114. The partial pressures can be determined by using the mole fractions.

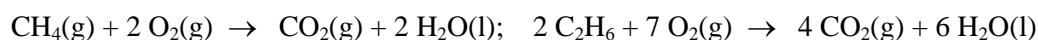
$$P_{\text{methane}} = P_{\text{total}} \times \chi_{\text{methane}} = 1.44 \text{ atm} \times 0.915 = 1.32 \text{ atm}; \quad P_{\text{ethane}} = 1.44 - 1.32 = 0.12 \text{ atm}$$

Determining the number of moles of natural gas combusted:

$$n_{\text{natural gas}} = \frac{PV}{RT} = \frac{1.44 \text{ atm} \times 15.00 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 293 \text{ K}} = 0.898 \text{ mol natural gas}$$

$$n_{\text{methane}} = n_{\text{natural gas}} \times \chi_{\text{methane}} = 0.898 \text{ mol} \times 0.915 = 0.822 \text{ mol methane}$$

$$n_{\text{ethane}} = 0.898 - 0.822 = 0.076 \text{ mol ethane}$$



$$0.822 \text{ mol CH}_4 \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol CH}_4} \times \frac{18.02 \text{ g H}_2\text{O}}{\text{mol H}_2\text{O}} = 29.6 \text{ g H}_2\text{O}$$

$$0.076 \text{ mol C}_2\text{H}_6 \times \frac{6 \text{ mol H}_2\text{O}}{2 \text{ mol C}_2\text{H}_6} \times \frac{18.02 \text{ g H}_2\text{O}}{\text{mol H}_2\text{O}} = 4.1 \text{ g H}_2\text{O}$$

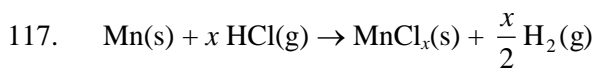
The total mass of H<sub>2</sub>O produced = 29.6 g + 4.1 g = 33.7 g H<sub>2</sub>O.

$$115. \quad P_{\text{total}} = P_{\text{H}_2} + P_{\text{H}_2\text{O}}, \quad 1.032 \text{ atm} = P_{\text{H}_2} + 32 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}, \quad 1.032 - 0.042 = 0.990 \text{ atm} = P_{\text{H}_2}$$

$$n_{\text{H}_2} = \frac{P_{\text{H}_2} V}{RT} = \frac{0.990 \text{ atm} \times 0.240 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 303 \text{ K}} = 9.56 \times 10^{-3} \text{ mol H}_2$$

$$9.56 \times 10^{-3} \text{ mol H}_2 \times \frac{1 \text{ mol Zn}}{\text{mol H}_2} \times \frac{65.38 \text{ g Zn}}{\text{mol Zn}} = 0.625 \text{ g Zn}$$

116.  $2 \text{HN}_3(\text{g}) \rightarrow 3 \text{N}_2(\text{g}) + \text{H}_2(\text{g})$ ; at constant V and T, P is directly proportional to n. In the reaction, we go from 2 moles of gaseous reactants to 4 moles of gaseous products. Because moles doubled, the final pressure will double ( $P_{\text{total}} = 6.0 \text{ atm}$ ). Similarly, from the 2 : 1 mole ratio between HN<sub>3</sub> and H<sub>2</sub>, the partial pressure of H<sub>2</sub> will be  $3.0/2 = 1.5 \text{ atm}$ . The partial pressure of N<sub>2</sub> will be  $(3/2)3.0 \text{ atm} = 4.5 \text{ atm}$ . This is from the 2 : 3 mole ratio between HN<sub>3</sub> and N<sub>2</sub>.



$$n_{\text{H}_2} = \frac{PV}{RT} = \frac{0.951 \text{ atm} \times 3.22 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 373 \text{ K}} = 0.100 \text{ mol H}_2$$

$$\text{Mol Cl in compound} = \text{mol HCl} = 0.100 \text{ mol H}_2 \times \frac{x \text{ mol Cl}}{\frac{x}{2} \text{ mol H}_2} = 0.200 \text{ mol Cl}$$

$$\frac{\text{Mol Cl}}{\text{Mol Mn}} = \frac{0.200 \text{ mol Cl}}{2.747 \text{ g Mn} \times \frac{1 \text{ mol Mn}}{54.94 \text{ g Mn}}} = \frac{0.200 \text{ mol Cl}}{0.05000 \text{ mol Mn}} = 4.00$$

The formula of compound is  $\text{MnCl}_4$ .

$$118. \quad \text{a. Volume of hot air: } V = \frac{4}{3} \pi r^3 = \frac{4}{3} \pi (2.50 \text{ m})^3 = 65.4 \text{ m}^3$$

(Note: Radius = diameter/2 = 5.00/2 = 2.50 m)

$$65.4 \text{ m}^3 \times \left( \frac{10 \text{ dm}}{\text{m}} \right)^3 \times \frac{1 \text{ L}}{\text{dm}^3} = 6.54 \times 10^4 \text{ L}$$

$$n = \frac{PV}{RT} = \frac{\left( 745 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} \right) \times 6.54 \times 10^4 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times (273 + 65) \text{ K}} = 2.31 \times 10^3 \text{ mol air}$$

$$\text{Mass of hot air} = 2.31 \times 10^3 \text{ mol} \times \frac{29.0 \text{ g}}{\text{mol}} = 6.70 \times 10^4 \text{ g}$$

$$\text{Air displaced: } n = \frac{PV}{RT} = \frac{\frac{745}{760} \text{ atm} \times 6.54 \times 10^4 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times (273 + 21) \text{ K}} = 2.66 \times 10^3 \text{ mol air}$$

$$\text{Mass of air displaced} = 2.66 \times 10^3 \text{ mol} \times \frac{29.0 \text{ g}}{\text{mol}} = 7.71 \times 10^4 \text{ g}$$

$$\text{Lift} = 7.71 \times 10^4 \text{ g} - 6.70 \times 10^4 \text{ g} = 1.01 \times 10^4 \text{ g}$$



- b. Mass of air displaced is the same,  $7.71 \times 10^4$  g. Moles of He in balloon will be the same as moles of air displaced,  $2.66 \times 10^3$  mol, because P, V, and T are the same.

$$\text{Mass of He} = 2.66 \times 10^3 \text{ mol} \times \frac{4.003 \text{ g}}{\text{mol}} = 1.06 \times 10^4 \text{ g}$$

$$\text{Lift} = 7.71 \times 10^4 \text{ g} - 1.06 \times 10^4 \text{ g} = 6.65 \times 10^4 \text{ g}$$

c. Hot air:  $n = \frac{PV}{RT} = \frac{\frac{630.}{760} \text{ atm} \times (6.54 \times 10^4 \text{ L})}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 338 \text{ K}} = 1.95 \times 10^3 \text{ mol air}$

$$1.95 \times 10^3 \text{ mol} \times \frac{29.0 \text{ g}}{\text{mol}} = 5.66 \times 10^4 \text{ g of hot air}$$

Air displaced:  $n = \frac{PV}{RT} = \frac{\frac{630.}{760} \text{ atm} \times (6.54 \times 10^4 \text{ L})}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 294 \text{ K}} = 2.25 \times 10^3 \text{ mol air}$

$$2.25 \times 10^3 \text{ mol} \times \frac{29.0 \text{ g}}{\text{mol}} = 6.53 \times 10^4 \text{ g of air displaced}$$

$$\text{Lift} = 6.53 \times 10^4 \text{ g} - 5.66 \times 10^4 \text{ g} = 8.7 \times 10^3 \text{ g}$$

- d. Mass of hot air =  $6.70 \times 10^4$  g (from part a)

Air displaced:  $n = \frac{PV}{RT} = \frac{\frac{745}{760} \text{ atm} \times (6.54 \times 10^4 \text{ L})}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 265 \text{ K}} = 2.95 \times 10^3 \text{ mol air}$

$$2.95 \times 10^3 \text{ mol} \times \frac{29.0 \text{ g}}{\text{mol}} = 8.56 \times 10^4 \text{ g of air displaced}$$

$$\text{Lift} = 8.56 \times 10^4 \text{ g} - 6.70 \times 10^4 \text{ g} = 1.86 \times 10^4 \text{ g}$$

119.  $PV = nRT$ , V and T are constant.  $\frac{P_1}{n_1} = \frac{P_2}{n_2}$  or  $\frac{P_1}{P_2} = \frac{n_1}{n_2}$

When V and T are constant, then pressure is directly proportional to moles of gas present, and pressure ratios are identical to mole ratios.

At  $25^\circ\text{C}$ :  $2 \text{ H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{ H}_2\text{O}(\text{l})$ ,  $\text{H}_2\text{O}(\text{l})$  is produced at  $25^\circ\text{C}$ .

The balanced equation requires 2 mol H<sub>2</sub> for every mol O<sub>2</sub> reacted. The same ratio (2 : 1) holds true for pressure units. The actual pressure ratio present is 2 atm H<sub>2</sub> to 3 atm O<sub>2</sub>, well below the required 2 : 1 ratio. Therefore, H<sub>2</sub> is the limiting reagent. The only gas present at 25°C after the reaction goes to completion will be the excess O<sub>2</sub>.

$$P_{\text{O}_2}(\text{reacted}) = 2.00 \text{ atm H}_2 \times \frac{1 \text{ atm O}_2}{2 \text{ atm H}_2} = 1.00 \text{ atm O}_2$$

$$P_{\text{O}_2}(\text{excess}) = P_{\text{O}_2}(\text{initial}) - P_{\text{O}_2}(\text{reacted}) = 3.00 \text{ atm} - 1.00 \text{ atm} = 2.00 \text{ atm O}_2 = P_{\text{total}}$$

At 125°C: 2 H<sub>2</sub>(g) + O<sub>2</sub>(g) → 2 H<sub>2</sub>O(g), H<sub>2</sub>O(g) is produced at 125°C.

The major difference in the problem is that gaseous H<sub>2</sub>O is now a product (instead of liquid H<sub>2</sub>O), which will increase the total pressure because an additional gas is present.

$$P_{\text{H}_2\text{O}}(\text{produced}) = 2.00 \text{ atm H}_2 \times \frac{2 \text{ atm H}_2\text{O}}{2 \text{ atm H}_2} = 2.00 \text{ atm H}_2\text{O}$$

$$P_{\text{total}} = P_{\text{O}_2}(\text{excess}) + P_{\text{H}_2\text{O}}(\text{produced}) = 2.00 \text{ atm O}_2 + 2.00 \text{ atm H}_2\text{O} = 4.00 \text{ atm}$$

120. Average velocity  $\propto (1/M)^{1/2}$  at constant T; the pressure in container A will increase initially because the lighter H<sub>2</sub> molecules will effuse into container A faster than air will escape container A. However, the pressures will eventually equalize once the gases have had time to mix thoroughly.

$$121. \text{ Mol of He removed} = \frac{PV}{RT} = \frac{1.00 \text{ atm} \times (1.75 \times 10^{-3} \text{ L})}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298 \text{ K}} = 7.16 \times 10^{-5} \text{ mol He}$$

In the original flask,  $7.16 \times 10^{-5}$  mol of He exerted a partial pressure of 1.960 – 1.710 = 0.250 atm.

$$V = \frac{nRT}{P} = \frac{(7.16 \times 10^{-5} \text{ mol}) \times 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 298 \text{ K}}{0.250 \text{ atm}} = 7.00 \times 10^{-3} \text{ L} \\ = 7.00 \text{ mL}$$

122. a. Initially  $P_{\text{N}_2} = P_{\text{H}_2} = 1.00 \text{ atm}$  and the total pressure is 2.00 atm ( $P_{\text{total}} = P_{\text{N}_2} + P_{\text{H}_2}$ ). The total pressure after reaction will also be 2.00 atm because we have a constant-pressure container. Because V and T are constant before the reaction takes place, there must be equal moles of N<sub>2</sub> and H<sub>2</sub> present initially. Let  $x = \text{mol N}_2 = \text{mol H}_2$  that are present initially. From the balanced equation, N<sub>2</sub>(g) + 3 H<sub>2</sub>(g) → 2 NH<sub>3</sub>(g), H<sub>2</sub> will be limiting because three times as many moles of H<sub>2</sub> are required to react as compared to moles of N<sub>2</sub>.

After the reaction occurs, none of the  $\text{H}_2$  remains (it is the limiting reagent).

$$\text{Mol NH}_3 \text{ produced} = x \text{ mol H}_2 \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} = 2x/3$$

$$\text{Mol N}_2 \text{ reacted} = x \text{ mol H}_2 \times \frac{1 \text{ mol N}_2}{3 \text{ mol H}_2} = x/3$$

$$\text{Mol N}_2 \text{ remaining} = x \text{ mol N}_2 \text{ present initially} - x/3 \text{ mol N}_2 \text{ reacted} = 2x/3 \text{ mol N}_2$$

After the reaction goes to completion, equal moles of  $\text{N}_2(\text{g})$  and  $\text{NH}_3(\text{g})$  are present ( $2x/3$ ). Because equal moles are present, the partial pressure of each gas must be equal ( $P_{\text{N}_2} = P_{\text{NH}_3}$ ).

$$P_{\text{total}} = 2.00 \text{ atm} = P_{\text{N}_2} + P_{\text{NH}_3}; \text{ solving: } P_{\text{N}_2} = 1.00 \text{ atm} = P_{\text{NH}_3}$$

- b.  $V \propto n$  because  $P$  and  $T$  are constant. The moles of gas present initially are:

$$n_{\text{N}_2} + n_{\text{H}_2} = x + x = 2x \text{ mol}$$

After reaction, the moles of gas present are:

$$n_{\text{N}_2} + n_{\text{NH}_3} = \frac{2x}{3} + \frac{2x}{3} = 4x/3 \text{ mol}$$

$$\frac{V_{\text{after}}}{V_{\text{initial}}} = \frac{n_{\text{after}}}{n_{\text{initial}}} = \frac{4x/3}{2x} = \frac{2}{3}$$

The volume of the container will be two-thirds the original volume, so:

$$V = 2/3(15.0 \text{ L}) = 10.0 \text{ L}$$

123. a.  $2 \text{ CH}_4(\text{g}) + 2 \text{ NH}_3(\text{g}) + 3 \text{ O}_2(\text{g}) \rightarrow 2 \text{ HCN}(\text{g}) + 6 \text{ H}_2\text{O}(\text{g})$
- b. Volumes of gases are proportional to moles at constant  $T$  and  $P$ . Using the balanced equation, methane and ammonia are in stoichiometric amounts and oxygen is in excess. In 1 second:

$$n_{\text{CH}_4} = \frac{PV}{RT} = \frac{1.00 \text{ atm} \times 20.0 \text{ L}}{0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 423 \text{ K}} = 0.576 \text{ mol CH}_4$$

$$\frac{0.576 \text{ mol CH}_4}{\text{s}} \times \frac{2 \text{ mol HCN}}{2 \text{ mol CH}_4} \times \frac{27.03 \text{ g HCN}}{\text{mol HCN}} = 15.6 \text{ g HCN/s}$$

124. a. Out of 100.00 g of Z, we have:

$$34.38 \text{ g Ni} \times \frac{1 \text{ mol}}{58.69 \text{ g}} = 0.5858 \text{ mol Ni}$$

$$28.13 \text{ g C} \times \frac{1 \text{ mol}}{12.011 \text{ g}} = 2.342 \text{ mol C}; \quad \frac{2.342}{0.5858} = 3.998$$

$$37.48 \text{ g O} \times \frac{1 \text{ mol}}{15.999 \text{ g}} = 2.343 \text{ mol O}; \quad \frac{2.343}{0.5858} = 4.000$$

The empirical formula is  $\text{NiC}_4\text{O}_4$ .

b.  $\frac{\text{Rate Z}}{\text{Rate Ar}} = \left( \frac{M_{\text{Ar}}}{M_{\text{Z}}} \right)^{1/2} = \left( \frac{39.95}{M_{\text{Z}}} \right)^{1/2}$ ; because initial mol Ar = mol Z:

$$0.4837 = \left( \frac{39.95}{M_{\text{Z}}} \right)^{1/2}, \quad M_{\text{Z}} = 170.8 \text{ g/mol}$$

- c.  $\text{NiC}_4\text{O}_4$ :  $M = 58.69 + 4(12.01) + 4(16.00) = 170.73 \text{ g/mol}$

Molecular formula is also  $\text{NiC}_4\text{O}_4$ .

- d. Each effusion step changes the concentration of Z in the gas by a factor of 0.4837. The original concentration of Z molecules to Ar atoms is a 1 : 1 ratio. After 5 stages:

$$n_{\text{Z}}/n_{\text{Ar}} = (0.4837)^5 = 2.648 \times 10^{-2}$$

125.  $n_{\text{Ar}} = \frac{228 \text{ g}}{39.95 \text{ g/mol}} = 5.71 \text{ mol Ar}$ ;  $\chi_{\text{CH}_4} = \frac{n_{\text{CH}_4}}{n_{\text{CH}_4} + n_{\text{Ar}}}$ ,  $0.650 = \frac{n_{\text{CH}_4}}{n_{\text{CH}_4} + 5.71}$

$$0.650(n_{\text{CH}_4} + 5.71) = n_{\text{CH}_4}, \quad 3.71 = (0.350)n_{\text{CH}_4}, \quad n_{\text{CH}_4} = 10.6 \text{ mol CH}_4$$

$$\text{KE}_{\text{avg}} = \frac{3}{2} RT \text{ for 1 mol}$$

$$\text{Thus } \text{KE}_{\text{total}} = (10.6 + 5.71 \text{ mol}) \times 3/2 \times 8.3145 \text{ J K}^{-1} \text{ mol}^{-1} \times 298 \text{ K} = 6.06 \times 10^4 \text{ J} = 60.6 \text{ kJ}$$

126.  $2 \text{ SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{ SO}_3(\text{g})$ ; because P and T are constant, volume ratios will equal mole ratios ( $V_f/V_i = n_f/n_i$ ). Let  $x = \text{mol SO}_2 = \text{mol O}_2$  present initially.  $\text{SO}_2$  will be limiting because a 2 : 1  $\text{SO}_2$  to  $\text{O}_2$  mole ratio is required by the balanced equation, but only a 1 : 1 mole ratio is present. Therefore, no  $\text{SO}_2$  will be present after the reaction goes to completion. However, excess  $\text{O}_2(\text{g})$  will be present as well as the  $\text{SO}_3(\text{g})$  produced.

$$\text{Mol O}_2 \text{ reacted} = x \text{ mol SO}_2 \times \frac{1 \text{ mol O}_2}{2 \text{ mol SO}_2} = x/2 \text{ mol O}_2$$

$$\text{Mol O}_2 \text{ remaining} = x \text{ mol O}_2 \text{ initially} - x/2 \text{ mol O}_2 \text{ reacted} = x/2 \text{ mol O}_2$$

$$\text{Mol SO}_3 \text{ produced} = x \text{ mol SO}_2 \times \frac{2 \text{ mol SO}_3}{2 \text{ mol SO}_2} = x \text{ mol SO}_3$$

$$\text{Total moles gas initially} = x \text{ mol SO}_2 + x \text{ mol O}_2 = 2x$$

$$\text{Total moles gas after reaction} = x/2 \text{ mol O}_2 + x \text{ mol SO}_3 = (3/2)x = (1.5)x$$

$$\frac{n_f}{n_i} = \frac{V_f}{V_i} = \frac{(1.5)x}{2x} = \frac{1.5}{2} = 0.75; \quad V_f/V_i = 0.75 : 1 \text{ or } 3 : 4$$

The volume of the reaction container shrinks to 75% of the initial volume.

127.  $P_1V_1 = P_2V_2$ ; the total volume is 1.00 L + 1.00 L + 2.00 L = 4.00 L.

$$\text{For He: } P_2 = \frac{P_1V_1}{V_2} = 200. \text{ torr} \times \frac{1.00 \text{ L}}{4.00 \text{ L}} = 50.0 \text{ torr He}$$

$$\text{For Ne: } P_2 = 0.400 \text{ atm} \times \frac{1.00 \text{ L}}{4.00 \text{ L}} = 0.100 \text{ atm}; \quad 0.100 \text{ atm} \times \frac{760 \text{ torr}}{\text{atm}} = 76.0 \text{ torr Ne}$$

$$\text{For Ar: } P_2 = 24.0 \text{ kPa} \times \frac{2.00 \text{ L}}{4.00 \text{ L}} = 12.0 \text{ kPa}; \quad 12.0 \text{ kPa} \times \frac{1 \text{ atm}}{101.3 \text{ kPa}} \times \frac{760 \text{ torr}}{\text{atm}} = 90.0 \text{ torr Ar}$$

$$P_{\text{total}} = 50.0 + 76.0 + 90.0 = 216.0 \text{ torr}$$

128.  $33.5 \text{ mg CO}_2 \times \frac{12.01 \text{ mg C}}{44.01 \text{ mg CO}_2} = 9.14 \text{ mg C}$ ;  $\% \text{ C} = \frac{9.14 \text{ mg}}{35.0 \text{ mg}} \times 100 = 26.1\% \text{ C}$

$$41.1 \text{ mg H}_2\text{O} \times \frac{2.016 \text{ mg H}}{18.02 \text{ mg H}_2\text{O}} = 4.60 \text{ mg H}$$
;  $\% \text{ H} = \frac{4.60 \text{ mg}}{35.0 \text{ mg}} \times 100 = 13.1\% \text{ H}$

$$n_{\text{N}_2} = \frac{P_{\text{N}_2} V}{RT} = \frac{740. \text{ atm} \times 35.6 \times 10^{-3} \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298 \text{ K}} = 1.42 \times 10^{-3} \text{ mol N}_2$$

$$1.42 \times 10^{-3} \text{ mol N}_2 \times \frac{28.02 \text{ g N}_2}{\text{mol N}_2} = 3.98 \times 10^{-2} \text{ g nitrogen} = 39.8 \text{ mg nitrogen}$$

$$\text{Mass \% N} = \frac{39.8 \text{ mg}}{65.2 \text{ mg}} \times 100 = 61.0\% \text{ N}$$

Or we can get % N by difference:  $\% \text{ N} = 100.0 - (26.1 + 13.1) = 60.8\%$

Out of 100.0 g:

$$26.1 \text{ g C} \times \frac{1 \text{ mol}}{12.01 \text{ g}} = 2.17 \text{ mol C}; \quad \frac{2.17}{2.17} = 1.00$$

$$13.1 \text{ g H} \times \frac{1 \text{ mol}}{1.008 \text{ g}} = 13.0 \text{ mol H}; \quad \frac{13.0}{2.17} = 5.99$$

$$60.8 \text{ g N} \times \frac{1 \text{ mol}}{14.01 \text{ g}} = 4.34 \text{ mol N}; \quad \frac{4.34}{2.17} = 2.00$$

Empirical formula is  $\text{CH}_6\text{N}_2$ .

$$\frac{\text{Rate}_1}{\text{Rate}_2} = \left( \frac{M}{39.95} \right)^{1/2} = \frac{26.4}{24.6} = 1.07, \quad M = (1.07)^2 \times 39.95 = 45.7 \text{ g/mol}$$

Empirical formula mass of  $\text{CH}_6\text{N}_2 \approx 12 + 6 + 28 = 46$ . Thus the molecular formula is also  $\text{CH}_6\text{N}_2$ .

129. a.  $156 \text{ mL} \times \frac{1.34 \text{ g}}{\text{mL}} = 209 \text{ g HSiCl}_3 = \text{actual yield of HSiCl}_3$

$$n_{\text{HCl}} = \frac{PV}{RT} = \frac{10.0 \text{ atm} \times 15.0 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 308 \text{ K}} = 5.93 \text{ mol HCl}$$

$$5.93 \text{ mol HCl} \times \frac{1 \text{ mol HSiCl}_3}{3 \text{ mol HCl}} \times \frac{135.45 \text{ g HSiCl}_3}{\text{mol HSiCl}_3} = 268 \text{ g HSiCl}_3$$

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{209 \text{ g}}{268 \text{ g}} \times 100 = 78.0\%$$

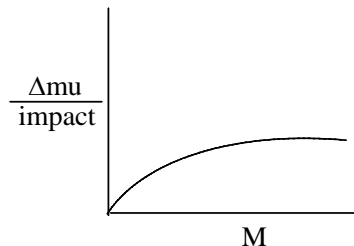
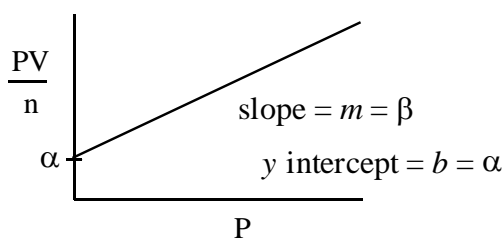
b.  $209 \text{ g HSiCl}_3 \times \frac{1 \text{ mol HSiCl}_3}{135.45 \text{ g HSiCl}_3} \times \frac{1 \text{ mol SiH}_4}{4 \text{ mol HSiCl}_3} = 0.386 \text{ mol SiH}_4$

This is the theoretical yield. If the percent yield is 93.1%, then the actual yield is:

$$0.386 \text{ mol SiH}_4 \times 0.931 = 0.359 \text{ mol SiH}_4$$

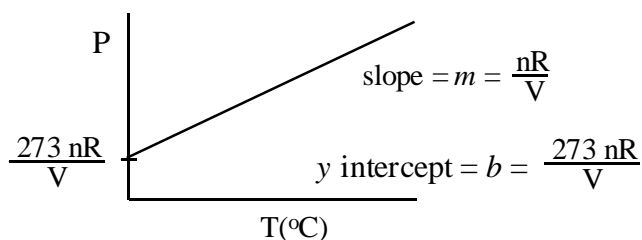
$$V_{\text{SiH}_4} = \frac{nRT}{P} = \frac{0.359 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 308 \text{ K}}{10.0 \text{ atm}} = 0.907 \text{ L} = 907 \text{ mL SiH}_4$$

130. a.  $\frac{PV}{n} = \alpha + \beta P$  (straight line,  $y = b + mx$ )
- b.  $\frac{\Delta(\mu)}{\text{impact}} = 2\mu \propto M \left(\frac{T}{M}\right)^{1/2} = \sqrt{M}$  at constant T



c.  $T_K = T_{0C} + 273$ ;  $P = \frac{nR(T_{0C} + 273)}{V} = \text{constant}(T_{0C} + 273)$ , where  $\text{constant} = \frac{nR}{V}$ .

This is in the form of the straight line equation,  $y = mx + b$ .



131. We will apply Boyle's law to solve.  $PV = nRT = \text{constant}$ ,  $P_1V_1 = P_2V_2$

Let condition (1) correspond to He from the tank that can be used to fill balloons. We must leave 1.0 atm of He in the tank, so  $P_1 = 200. - 1.00 = 199$  atm and  $V_1 = 15.0$  L. Condition (2) will correspond to the filled balloons with  $P_2 = 1.00$  atm and  $V_2 = N(2.00$  L), where  $N$  is the number of filled balloons, each at a volume of 2.00 L.

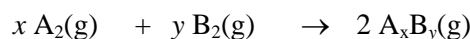
$199 \text{ atm} \times 15.0 \text{ L} = 1.00 \text{ atm} \times N(2.00 \text{ L})$ ,  $N = 1492.5$ ; we can't fill 0.5 of a balloon, so  $N = 1492$  balloons, or to 3 significant figures, 1490 balloons.

### Challenge Problems

132. a. The formula of the compound  $A_xB_y$  depends on which gas is limiting,  $A_2$  or  $B_2$ . We need to determine both possible products. The procedure we will use is to assume one reactant is limiting, and then determine what happens to the initial total moles of gas as it is converted into the product. Because  $P$  and  $T$  are constant,  $\text{volume} \propto n$ . Because mass is conserved in a chemical reaction, any change in density must be due to a change in volume of the container as the reaction goes to completion.

$$\text{Density} = d \propto \frac{1}{V} \quad \text{and} \quad V \propto n, \quad \text{so:} \quad \frac{d_{\text{after}}}{d_{\text{initial}}} = \frac{n_{\text{initial}}}{n_{\text{after}}}$$

Assume the molecular formula of the product is  $A_xB_y$ , where  $x$  and  $y$  are whole numbers. First, let's consider when  $A_2$  is limiting with  $x$  moles each of  $A_2$  and  $B_2$  in our equimolar mixture. Note that the coefficient in front of  $A_xB_y$  in the equation must be 2 for a balanced reaction.



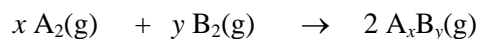
Initial	$x$ mol	$x$ mol	0 mol
Change	$-x$ mol	$-y$ mol	$+2$ mol
Final	0	$(x - y)$ mol	2 mol

$$\frac{d_{\text{after}}}{d_{\text{initial}}} = 1.50 = \frac{n_{\text{initial}}}{n_{\text{after}}} = \frac{2x}{x - y + 2}$$

$$(1.50)x - (1.50)y + 3.00 = 2x, \quad 3.00 - (1.50)y = (0.50)x$$

Because  $x$  and  $y$  are whole numbers,  $y$  must be 1 because the above equation does not allow  $y$  to be 2 or greater. When  $y = 1$ ,  $x = 3$  giving a formula of  $A_3B$  if  $A_2$  is limiting.

Assuming  $B_2$  is limiting with  $y$  moles in the equimolar mixture:

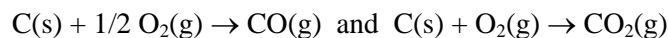


Initial	$y$	$y$	0
Change	$-x$	$-y$	$+2$
After	$y - x$	0	2

$$\frac{\text{density}_{\text{after}}}{\text{density}_{\text{before}}} = 1.50 = \frac{n_{\text{initial}}}{n_{\text{after}}} = \frac{2y}{y - x + 2}$$

Solving gives  $x = 1$  and  $y = 3$  for a molecular formula of  $AB_3$  when  $B_2$  is limiting.

- b. In both possible products, the equations dictated that only one mole of either A or B had to be present in the formula. Any number larger than 1 would not fit the data given in the problem. Thus the two formulas determined are both molecular formulas and not just empirical formulas.
133. The reactions are:



$$PV = nRT, \quad P = n \left( \frac{RT}{V} \right) = n(\text{constant})$$

Because the pressure has increased by 17.0%, the number of moles of gas has also increased by 17.0%.

$$n_{\text{final}} = (1.170)n_{\text{initial}} = 1.170(5.00) = 5.85 \text{ mol gas} = n_{O_2} + n_{CO} + n_{CO_2}$$



$n_{\text{CO}} + n_{\text{CO}_2} = 5.00$  (balancing moles of C). Solving by simultaneous equations:

$$\begin{array}{r} n_{\text{O}_2} + n_{\text{CO}} + n_{\text{CO}_2} = 5.85 \\ -(n_{\text{CO}} + n_{\text{CO}_2} = 5.00) \\ \hline n_{\text{O}_2} = 0.85 \end{array}$$

If all C were converted to  $\text{CO}_2$ , no  $\text{O}_2$  would be left. If all C were converted to CO, we would get 5 mol CO and 2.5 mol excess  $\text{O}_2$  in the reaction mixture. In the final mixture, moles of CO equals twice the moles of  $\text{O}_2$  present ( $n_{\text{CO}} = 2n_{\text{O}_2}$ ).

$$n_{\text{CO}} = 2n_{\text{O}_2} = 1.70 \text{ mol CO}; 1.70 + n_{\text{CO}_2} = 5.00, n_{\text{CO}_2} = 3.30 \text{ mol CO}_2$$

$$\chi_{\text{CO}} = \frac{1.70}{5.85} = 0.291; \chi_{\text{CO}_2} = \frac{3.30}{5.85} = 0.564; \chi_{\text{O}_2} = 1.000 - 0.291 - 0.564 = 0.145$$

134.  $\text{BaO(s)} + \text{CO}_2\text{(g)} \rightarrow \text{BaCO}_3\text{(s)}$ ;  $\text{CaO(s)} + \text{CO}_2\text{(g)} \rightarrow \text{CaCO}_3\text{(s)}$

$$n_i = \frac{P_i V}{RT} = \text{initial moles of CO}_2 = \frac{\frac{750.}{760} \text{ atm} \times 1.50 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 303.2 \text{ K}} = 0.0595 \text{ mol CO}_2$$

$$n_f = \frac{P_f V}{RT} = \text{final moles of CO}_2 = \frac{\frac{230.}{760} \text{ atm} \times 1.50 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 303.2 \text{ K}} = 0.0182 \text{ mol CO}_2$$

$$0.0595 - 0.0182 = 0.0413 \text{ mol CO}_2 \text{ reacted}$$

Because each metal reacts 1 : 1 with  $\text{CO}_2$ , the mixture contains 0.0413 mol of BaO and CaO. The molar masses of BaO and CaO are 153.3 and 56.08 g/mol, respectively.

Let  $x$  = mass of BaO and  $y$  = mass of CaO, so:

$$x + y = 5.14 \text{ g and } \frac{x}{153.3} + \frac{y}{56.08} = 0.0413 \text{ mol or } x + (2.734)y = 6.33$$

Solving by simultaneous equations:

$$\begin{array}{r} x + (2.734)y = 6.33 \\ -x \quad \quad -y = -5.14 \\ \hline (1.734)y = 1.19, y = 1.19/1.734 = 0.686 \end{array}$$

$$y = 0.686 \text{ g CaO and } 5.14 - y = x = 4.45 \text{ g BaO}$$

$$\text{Mass \% BaO} = \frac{4.45 \text{ g BaO}}{5.14 \text{ g}} \times 100 = 86.6\% \text{ BaO}$$

$$\% \text{ CaO} = 100.0 - 86.6 = 13.4\% \text{ CaO}$$

135. Figure 5.16 shows the effect of temperature on the Maxwell-Boltzmann distribution of velocities of molecules. Note that as temperature increases, the probability that a gas particle has the most probable velocity decreases. Thus, since the probability of the gas particle with the most probable velocity decreased by one-half, then the temperature must be higher than 300. K.

The equation that determines the probability that a gas molecule has a certain velocity is:

$$f(u) = 4\pi \left( \frac{m}{2\pi k_B T} \right)^{3/2} u^2 e^{-mu^2/2k_B T}$$

Let  $T_x$  = the unknown temperature, then:

$$\frac{f(u_{mp,x})}{f(u_{mp,300})} = \frac{1}{2} = \frac{4\pi \left( \frac{m}{2\pi k_B T_x} \right)^{3/2} u_{mp,x}^2 e^{-mu_{mp,x}^2/2k_B T_x}}{4\pi \left( \frac{m}{2\pi k_B T_{300}} \right)^{3/2} u_{mp,300}^2 e^{-mu_{mp,300}^2/2k_B T_{300}}}$$

Because  $u_{mp} = \sqrt{\frac{2k_B T}{m}}$ , the equation reduces to:

$$\frac{1}{2} = \frac{\left( \frac{1}{T_x} \right)^{3/2} (T_x)}{\left( \frac{1}{T_{300}} \right)^{3/2} (T_{300})} = \left( \frac{T_{300}}{T_x} \right)^{1/2}$$

Note that the overall exponent term cancels from the expression when  $2k_B T/m$  is substituted for  $u_{mp}^2$  in the exponent term; the temperatures cancel. Solving for  $T_x$ :

$$\frac{1}{2} = \left( \frac{300. \text{ K}}{T_x} \right)^{1/2}, \quad T_x = 1.20 \times 10^3 \text{ K}$$

As expected,  $T_x$  is higher than 300. K.

136. a. The number of collisions of gas particles with the walls of the container is proportional to:

$$Z_A \propto \frac{N}{V} \sqrt{\frac{T}{M}}$$

where  $N$  = number of gas particles,  $V$  = volume of container,  $T$  = temperature (Kelvin), and  $M$  = molar mass of gas particles in kilograms. Because both He samples are in separate containers of the same volume,  $V$  and  $M$  are constant. Because pressure and volume are constant,  $P \propto nT$  (also,  $n \propto N$ ).

Thus:  $Z_A \propto N\sqrt{T}$

$$\frac{Z_1}{Z_2} = \frac{N_1\sqrt{T_1}}{N_2\sqrt{T_2}} = 2, \text{ and } N_1T_1 = N_2T_2$$

$$\text{Thus: } \frac{N_1}{N_2} = \frac{2\sqrt{T_2}}{\sqrt{T_1}} = \frac{T_2}{T_1}, \quad \frac{2T_1}{\sqrt{T_1}} = \frac{T_2}{\sqrt{T_2}}, \quad 2\sqrt{T_1} = \sqrt{T_2}$$

Solving:  $4T_1 = T_2$ ,  $T_1 = 1/4 T_2$ ; because  $P \propto nT$ , and  $P$  is constant,  $n_1 = 4n_2$ .

Although the number of collisions in container 1 is twice as high, the temperature is one-fourth that of container 2. This is so because there are four times the number of moles of helium gas in container 1.

- b. There are twice the number of collisions, but because the temperature is lower, the gas particles are hitting with less forceful collisions. Overall, the pressure is the same in each container.

137. From the problem, we want  $Z_A/Z = 1.00 \times 10^{18}$  where  $Z_A$  is the collision frequency of the gas particles with the walls of the container and  $Z$  is the intermolecular collision frequency.

$$\text{From the text: } \frac{Z_A}{Z} = \frac{A \frac{N}{V} \sqrt{\frac{RT}{2\pi M}}}{4 \frac{N}{V} d^2 \sqrt{\frac{\pi RT}{M}}} = 1.00 \times 10^{18}, \quad 1.00 \times 10^{18} = \frac{A}{4 d^2 \pi \sqrt{2}}$$

If  $l$  = length of the cube edge container, then the area  $A$  of one cube face is  $l^2$  and the total area in the cube is  $6l^2$  (6 faces/cube). He diameter =  $d = 2(3.2 \times 10^{-11} \text{ m}) = 6.4 \times 10^{-11} \text{ m}$ .

Solving the above expression for  $A$ , and then for  $l$  gives  $l = 0.11 \text{ m} = 1.1 \text{ dm}$ .

$$\text{Volume} = l^3 = (1.1 \text{ dm})^3 = 1.3 \text{ dm}^3 = 1.3 \text{ L}$$

138. a. We assumed a pressure of 1.0 atm and a temperature of 25°C (298 K).

$$50. \text{ lb} \times 0.454 \text{ kg/lb} = 23 \text{ kg}$$

$$n = \frac{PV}{RT} = \frac{1.0 \text{ atm} \times 10. \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298 \text{ K}} = 0.41 \text{ mol gas}$$

The lift of one balloon is:  $0.41 \text{ mol}(29 \text{ g/mol} - 4.003 \text{ g/mol}) = 10. \text{ g}$ .

To lift 23 kg = 23,000 g, we need at least  $23,000/10 = 2300$  balloons. This is a lot of balloons.

- b. The balloon displaces air as it is filled. The displaced air has mass, as does the helium in the balloon, but the displaced air has more mass than the helium. The difference in this mass is the lift of the balloon. Because volume is constant, the difference in mass is directly related to the difference in density between air and helium.

139. Molar mass =  $\frac{dRT}{P}$ ,  $P$  and molar mass are constant;  $dT = \frac{P \times \text{molar mass}}{R} = \text{constant}$

$d = \text{constant}(1/T)$  or  $d_1T_1 = d_2T_2$ , where  $T$  is in kelvin (K).

$$T = x + ^\circ\text{C}; 1.2930(x + 0.0) = 0.9460(x + 100.0)$$

$$(1.2930)x = (0.9460)x + 94.60, (0.3470)x = 94.60, x = 272.6$$

From these data, absolute zero would be  $-272.6^\circ\text{C}$ . The actual value is  $-273.15^\circ\text{C}$ .

140. Dalton's law states:  $P_{\text{total}} = P_1 + P_2 + \dots + P_k$ , for  $k$  different types of gas molecules in a mixture. The postulates of the kinetic molecular theory are:

1. the volume of the individual particles can be assumed to be negligible.
2. the collisions of the particles with the walls of the container are the cause of the pressure exerted by the gas.
3. the particles assert no forces on each other.
4. the average kinetic energy of a collection of gas particles is assumed to be directly proportional to the Kelvin temperature of the gas.

The derivation is very similar to the ideal gas law derivation covered in Section 5.6 of the text. For a mixture of gases in a cube, there exist  $k$  different types of gas molecules. For each type ( $i$ ) of gas molecule, the force on the cube =  $F_i = (2m_i/L)u_i^2$ .

Because the gas particles are assumed non-interacting, the total force for all the gas molecules in the mixture is:

$$F_{\text{total}} = \sum_{i=1}^k \frac{2m_i}{L}(u_i^2)$$

Now we want the average force for each type of gas particle, which is:

$$\bar{F}_{\text{total}} = \sum_{i=1}^k \frac{2m_i}{L}(\bar{u}_i^2)$$

Pressure due to the average particle in this gas mixture of  $k$  types is the average total force divided by the total area. The expression for pressure is:

$$P = \frac{\sum_{i=1}^k \frac{2m_i}{L} (\bar{u}_i^2)}{6L^2} = \frac{\sum_{i=1}^k m_i (\bar{u}_i^2)}{3V}, \text{ where } V \text{ is the volume of the cube}$$

Total pressure due to the number of moles of different gases is:

$$P_{\text{total}} = \frac{\sum_{i=1}^k n_i N_A m_i (\bar{u}_i^2)}{3V}, \text{ where } N_A = \text{Avogadro's number}$$

Because molar  $KE_{i, \text{avg}} = N_A \left( \frac{1}{2} m_i \bar{u}_i^2 \right)$ , the expression for total pressure can be written as:

$$P_{\text{total}} = \sum_{i=1}^k \frac{\frac{2}{3} n_i N_A \left( \frac{1}{2} m_i \bar{u}_i^2 \right)}{V} = \sum_{i=1}^k \frac{\frac{2}{3} n_i KE_{i, \text{avg}}}{V}$$

Assuming molar  $KE_{i, \text{avg}}$  is proportional to  $T$  and is equal to  $\frac{3}{2} RT$ , then:

$$P_{\text{total}} = \sum_{i=1}^k \frac{n_i RT}{V} = \sum_{i=1}^k P_i \text{ because } P_i = \frac{n_i RT}{V}$$

This is Dalton's law of partial pressure. Note that no additional assumptions are necessary other than the postulates of the kinetic molecular theory and the conclusions drawn from the ideal gas law derivation.

$$141. \quad \frac{PV}{nRT} = 1 + \beta P; \quad \frac{n}{V} \times \text{molar mass} = d$$

$$\frac{\text{molar mass}}{RT} \times \frac{P}{d} = 1 + \beta P, \quad \frac{P}{d} = \frac{RT}{\text{molar mass}} + \frac{\beta RTP}{\text{molar mass}}$$

This is in the equation for a straight line:  $y = b + mx$ . If we plot  $P/d$  versus  $P$  and extrapolate to  $P = 0$ , we get a  $y$  intercept =  $b = 1.398 = RT/\text{molar mass}$ .

$$\text{At } 0.00^\circ\text{C, molar mass} = \frac{0.08206 \times 273.15}{1.398} = 16.03 \text{ g/mol.}$$

142. a. When the balloon is heated, the balloon will expand ( $P$  and  $n$  remain constant). The mass of the balloon is the same, but the volume increases, so the density of the argon in the balloon decreases. When the density is less than that of air, the balloon will rise.

- b. Assuming the balloon has no mass, when the density of the argon equals the density of air, the balloon will float in air. Above this temperature, the balloon will rise.

$$d_{\text{air}} = \frac{P \cdot MM_{\text{air}}}{RT}, \text{ where } MM_{\text{air}} = \text{average molar mass of air}$$

$$MM_{\text{air}} = 0.790 \times 28.02 \text{ g/mol} + 0.210 \times 32.00 \text{ g/mol} = 28.9 \text{ g/mol}$$

$$d_{\text{air}} = \frac{1.00 \text{ atm} \times 28.9 \text{ g/mol}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298 \text{ K}} = 1.18 \text{ g/L}$$

$$d_{\text{argon}} = \frac{1.00 \text{ atm} \times 39.95 \text{ g/mol}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times T} = 1.18 \text{ g/L}, T = 413 \text{ K}$$

Heat the Ar above 413 K or 140.°C and the balloon will float.

143. Initially we have 1.00 mol CH<sub>4</sub> (16.0 g/mol = molar mass) and 2.00 mol O<sub>2</sub> (32.0 g/mol = molar mass).



$$b + c = 1.00 \text{ (C balance); } 2a = b + 2c + d \text{ (O balance)}$$

$$2d = 4 \text{ (H balance), } d = 2 = 2.00 \text{ mol H}_2\text{O}$$

$$V_{\text{initial}} = \frac{nRT}{P} = \frac{3.00 \text{ mol} \times 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 425 \text{ K}}{1.00 \text{ atm}} = 104.6 \text{ L (1 extra sig .fig.)}$$

$$\text{Density}_{\text{initial}} = \frac{80.0 \text{ g}}{104.6 \text{ L}} = 0.7648 \text{ g/L (1 extra significant figure)}$$

Because mass is constant:

$$\text{mass} = V_{\text{initial}} \times d_{\text{initial}} = V_{\text{final}} \times d_{\text{final}}, V_{\text{final}} = V_{\text{initial}} \times \frac{d_{\text{initial}}}{d_{\text{final}}} = 104.6 \text{ L} \times \frac{0.7648 \text{ g/L}}{0.7282 \text{ g/L}}$$

$$V_{\text{final}} = 109.9 \text{ L (1 extra significant figure)}$$

$$n_{\text{final}} = \frac{PV}{RT} = \frac{1.00 \text{ atm} \times 109.9 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 425 \text{ K}} = 3.15 \text{ total moles of gas}$$

Assuming an excess of O<sub>2</sub> is present after reaction, an expression for the total moles of gas present at completion is:

$$b + c + 2.00 + (2.00 - a) = 3.15; \text{ Note: } d = 2.00 \text{ mol H}_2\text{O was determined previously.}$$

Because  $b + c = 1.00$ , solving gives  $a = 1.85$  mol  $\text{O}_2$  reacted. Indeed,  $\text{O}_2$  is in excess.

From the O balance equation:

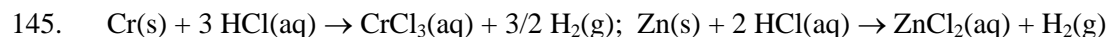
$$2a = 3.70 = b + 2c + 2.00, \quad b + 2c = 1.70$$

Because  $b + c = 1.00$ , solving gives  $b = 0.30$  mol  $\text{CO}$  and  $c = 0.70$  mol  $\text{CO}_2$ .

The fraction of methane that reacts to form  $\text{CO}$  is  $0.30 \text{ mol CO} / 1.00 \text{ mol CH}_4 = 0.30$  (or 30% by moles of the reacted methane forms  $\text{CO}$ ).

$$144. \quad Z_A = A \left( \frac{N}{V} \right) \left( \frac{RT}{2\pi M} \right)^{1/2}; \quad \frac{Z_1}{Z_2} = \frac{\left( \frac{T_1}{M_1} \right)^{1/2}}{\left( \frac{T_2}{M_2} \right)^{1/2}} = \left( \frac{M_2 T_1}{M_1 T_2} \right)^{1/2} = 1.00, \quad M_1 T_2 = M_2 T_1$$

$$\frac{T_2}{T_1} = \frac{M_2}{M_1}; \quad \frac{T_{\text{UF}_6}}{T_{\text{He}}} = \frac{M_{\text{UF}_6}}{M_{\text{He}}} = \frac{3520}{4.003} = 87.93$$



$$\text{Mol H}_2 \text{ produced} = n = \frac{PV}{RT} = \frac{\left( 750. \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} \right) \times 0.225 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times (273 + 27) \text{ K}} = 9.02 \times 10^{-3} \text{ mol H}_2$$

$$9.02 \times 10^{-3} \text{ mol H}_2 = \text{mol H}_2 \text{ from Cr reaction} + \text{mol H}_2 \text{ from Zn reaction}$$

$$\text{From the balanced equation: } 9.02 \times 10^{-3} \text{ mol H}_2 = \text{mol Cr} \times (3/2) + \text{mol Zn} \times 1$$

Let  $x$  = mass of Cr and  $y$  = mass of Zn, then:

$$x + y = 0.362 \text{ g and } 9.02 \times 10^{-3} = \frac{(1.5)x}{52.00} + \frac{y}{65.38}$$

We have two equations and two unknowns. Solving by simultaneous equations:

$$9.02 \times 10^{-3} = (0.02885)x + (0.01530)y$$

$$\underline{-0.01530 \times 0.362 = -(0.01530)x - (0.01530)y}$$

$$3.48 \times 10^{-3} = (0.01355)x, \quad x = \text{mass of Cr} = \frac{3.48 \times 10^{-3}}{0.01355} = 0.257 \text{ g}$$

$$y = \text{mass of Zn} = 0.362 \text{ g} - 0.257 \text{ g} = 0.105 \text{ g Zn}; \quad \text{mass \% Zn} = \frac{0.105 \text{ g}}{0.362 \text{ g}} \times 100 = 29.0\% \text{ Zn}$$

146. After the hole develops, assume each He that collides with the hole goes into the Rn side and that each Rn that collides with the hole goes into the He side. Assume no molecules return to the side in which they began. Initial moles of each gas:

$$n = \frac{PV}{RT} = \frac{(2.00 \times 10^{-6} \text{ atm}) \times 1.00 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 300. \text{ K}} = 8.12 \times 10^{-8} \text{ mol}$$

$$Z_{\text{He}} = A \times \frac{N}{V} \times \left( \frac{RT}{2\pi M} \right)^{1/2}, \quad \frac{N}{V} = \frac{P}{RT} \times N_A \times 1000 \text{ L/m}^3 \text{ and } A = \pi r^2$$

$$Z_{\text{He}} = \pi(1.00 \times 10^{-6} \text{ m})^2 \times \frac{2.00 \times 10^{-6}}{0.08206 \times 300.} \times (6.022 \times 10^{23}) \times 1000 \\ \times \left( \frac{8.3145 \times 300.}{2\pi(4.003 \times 10^{-3})} \right)^{1/2} = 4.84 \times 10^{10} \text{ collisions/s}$$

Therefore,  $4.84 \times 10^{10}$  atoms/s leave the He side.

$$10.0 \text{ h} \times \frac{60 \text{ min}}{1 \text{ h}} \times \frac{60 \text{ s}}{1 \text{ min}} \times \frac{4.84 \times 10^{10} \text{ atoms}}{\text{s}} = 1.74 \times 10^{15} \text{ atoms}$$

$$\text{or: } \frac{1.74 \times 10^{15} \text{ atoms}}{6.022 \times 10^{23} \text{ atoms/mol}} = 2.89 \times 10^{-9} \text{ mol He leave in 10.0 h.}$$

$$Z_{\text{Rn}} = \pi(1.00 \times 10^{-6} \text{ m})^2 \times \frac{2.00 \times 10^{-6}}{0.08206 \times 300.} \times (6.022 \times 10^{23}) \times 1000 \\ \times \left( \frac{8.3145 \times 300.}{2\pi(222 \times 10^{-3})} \right)^{1/2} = 6.50 \times 10^9 \text{ collisions/s}$$

$6.50 \times 10^9$  atoms/s leave Rn side.

$$3.60 \times 10^4 \text{ s} \times \frac{6.50 \times 10^9 \text{ atoms}}{\text{s}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}} = 3.89 \times 10^{-10} \text{ mol Rn leave in 10.0 h}$$

Side that began with He now contains:

$$8.12 \times 10^{-8} - 2.89 \times 10^{-9} = 7.83 \times 10^{-8} \text{ mol He} + 3.89 \times 10^{-10} \text{ mol Rn} \\ = 7.87 \times 10^{-8} \text{ moles total}$$

The pressure in the He side is:

$$P = \frac{nRT}{V} = \frac{(7.87 \times 10^{-8} \text{ mol}) \times 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 300. \text{ K}}{1.00 \text{ L}} = 1.94 \times 10^{-6} \text{ atm}$$



We can determine the pressure in the Rn chamber two ways. Because no gas has escaped, and because the initial pressures were equal and the pressure in one of the sides decreased by  $0.06 \times 10^{-6}$  atm, P in the second side must increase by  $0.06 \times 10^{-6}$  atm. Thus the pressure on the side that originally contained Rn is  $2.06 \times 10^{-6}$  atm. Or we can calculate P the same way as with He. The Rn side contains:

$$\begin{aligned} 8.12 \times 10^{-8} - 3.89 \times 10^{-10} &= 8.08 \times 10^{-8} \text{ mol Rn} + 2.89 \times 10^{-9} \text{ mol He} \\ &= 8.37 \times 10^{-8} \text{ mol total} \end{aligned}$$

$$P = \frac{nRT}{V} = \frac{(8.37 \times 10^{-8} \text{ mol}) \times 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 300. \text{ K}}{1.00 \text{ L}} = 2.06 \times 10^{-6} \text{ atm}$$

147. Each stage will give an enrichment of:

$$\frac{\text{Diffusion rate } ^{12}\text{CO}_2}{\text{Diffusion rate } ^{13}\text{CO}_2} = \left( \frac{M_{^{13}\text{CO}_2}}{M_{^{12}\text{CO}_2}} \right)^{1/2} = \left( \frac{45.001}{43.998} \right) = 1.0113$$

Because  $^{12}\text{CO}_2$  moves slightly faster, each successive stage will have less  $^{13}\text{CO}_2$ .

$$\frac{99.90 \text{ } ^{12}\text{CO}_2}{0.10 \text{ } ^{13}\text{CO}_2} \times 1.0113^N = \frac{99.990 \text{ } ^{12}\text{CO}_2}{0.010 \text{ } ^{13}\text{CO}_2}$$

$$1.0113^N = \frac{9,999.0}{999.00} = 10.009 \quad (\text{carrying extra significant figures})$$

$$N \log(1.0113) = \log(10.009), \quad N = \frac{1.000391}{4.88 \times 10^{-3}} = 2.05 \times 10^2 \approx 2.1 \times 10^2 \text{ stages are needed.}$$

148. Let  $x = \text{moles SO}_2 = \text{moles O}_2$  and  $z = \text{moles He}$ .

a.  $\frac{P \cdot MM}{RT}$  where MM = molar mass

$$1.924 \text{ g/L} = \frac{1.000 \text{ atm} \times MM}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 273.2 \text{ K}}, \quad MM_{\text{mixture}} = 43.13 \text{ g/mol}$$

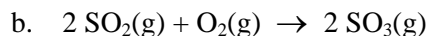
Assuming 1.000 total moles of mixture is present, then:  $x + x + z = 1.000$  and:

$$64.07 \text{ g/mol} \times x + 32.00 \text{ g/mol} \times x + 4.003 \text{ g/mol} \times z = 43.13 \text{ g}$$

$$2x + z = 1.000 \text{ and } (96.07)x + (4.003)z = 43.13$$

Solving:  $x = 0.4443 \text{ mol}$  and  $z = 0.1114 \text{ mol}$

Thus:  $\chi_{\text{He}} = 0.1114 \text{ mol}/1.000 \text{ mol} = 0.1114$



Initially, assume 0.4443 mol  $\text{SO}_2$ , 0.4443 mol  $\text{O}_2$  and 0.1114 mol He. Because  $\text{SO}_2$  is limiting, we end up with 0.2222 mol  $\text{O}_2$ , 0.4443 mol  $\text{SO}_3$ , and 0.1114 mol He in the gaseous product mixture. This gives:  $n_{\text{initial}} = 1.0000$  mol and  $n_{\text{final}} = 0.7779$  mol.

In a reaction, mass is constant.  $d = \frac{\text{mass}}{V}$  and  $V \propto n$  at constant P and T, so  $d \propto \frac{1}{n}$ .

$$\frac{n_{\text{initial}}}{n_{\text{final}}} = \frac{1.0000}{0.7779} = \frac{d_{\text{final}}}{d_{\text{initial}}}, \quad d_{\text{final}} = \left( \frac{1.0000}{0.7779} \right) \times 1.924 \text{ g/L}, \quad d_{\text{final}} = 2.473 \text{ g/L}$$

149.  $d = \text{molar mass}(P/RT)$ ; at constant P and T, the density of gas is directly proportional to the molar mass of the gas. Thus the molar mass of the gas has a value which is 1.38 times that of the molar mass of  $\text{O}_2$ .

$$\text{Molar mass} = 1.38(32.00 \text{ g/mol}) = 44.2 \text{ g/mol}$$

Because  $\text{H}_2\text{O}$  is produced when the unknown binary compound is combusted, the unknown must contain hydrogen. Let  $\text{A}_x\text{H}_y$  be the formula for unknown compound.

$$\text{Mol } \text{A}_x\text{H}_y = 10.0 \text{ g } \text{A}_x\text{H}_y \times \frac{1 \text{ mol } \text{A}_x\text{H}_y}{44.2 \text{ g}} = 0.226 \text{ mol } \text{A}_x\text{H}_y$$

$$\text{Mol H} = 16.3 \text{ g } \text{H}_2\text{O} \times \frac{1 \text{ mol } \text{H}_2\text{O}}{18.02 \text{ g}} \times \frac{2 \text{ mol H}}{\text{mol } \text{H}_2\text{O}} = 1.81 \text{ mol H}$$

$$\frac{1.81 \text{ mol H}}{0.226 \text{ mol } \text{A}_x\text{H}_y} = 8 \text{ mol H/mol } \text{A}_x\text{H}_y; \quad \text{A}_x\text{H}_y = \text{A}_x\text{H}_8$$

The mass of the  $x$  moles of A in the  $\text{A}_x\text{H}_8$  formula is:

$$44.2 \text{ g} - 8(1.008 \text{ g}) = 36.1 \text{ g}$$

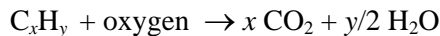
From the periodic table and by trial and error, some possibilities for  $\text{A}_x\text{H}_8$  are  $\text{ClH}_8$ ,  $\text{F}_2\text{H}_8$ ,  $\text{C}_3\text{H}_8$ , and  $\text{Be}_4\text{H}_8$ .  $\text{C}_3\text{H}_8$  and  $\text{Be}_4\text{H}_8$  fit the data best and because  $\text{C}_3\text{H}_8$  (propane) is a known substance,  $\text{C}_3\text{H}_8$  is the best possible identity from the data in this problem.

150. Assuming 1.000 L of the hydrocarbon ( $\text{C}_x\text{H}_y$ ), then the volume of products will be 4.000 L and the mass of products ( $\text{H}_2\text{O} + \text{CO}_2$ ) will be:

$$1.391 \text{ g/L} \times 4.000 \text{ L} = 5.564 \text{ g products}$$

$$\text{Mol } \text{C}_x\text{H}_y = n_{\text{C}_x\text{H}_y} = \frac{PV}{RT} = \frac{0.959 \text{ atm} \times 1.000 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298 \text{ K}} = 0.0392 \text{ mol}$$

$$\text{Mol products} = n_p = \frac{PV}{RT} = \frac{1.51 \text{ atm} \times 4.000 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 375 \text{ K}} = 0.196 \text{ mol}$$



Setting up two equations:

$$(0.0392)x + 0.0392(y/2) = 0.196 \text{ (moles of products)}$$

$$(0.0392)x(44.01 \text{ g/mol}) + 0.0392(y/2)(18.02 \text{ g/mol}) = 5.564 \text{ g (mass of products)}$$

Solving:  $x = 2$  and  $y = 6$ , so the formula of the hydrocarbon is  $\text{C}_2\text{H}_6$ .

151. a. The reaction is:  $\text{CH}_4(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$

$$PV = nRT, \quad \frac{PV}{n} = RT = \text{constant}, \quad \frac{P_{\text{CH}_4} V_{\text{CH}_4}}{n_{\text{CH}_4}} = \frac{P_{\text{air}} V_{\text{air}}}{n_{\text{air}}}$$

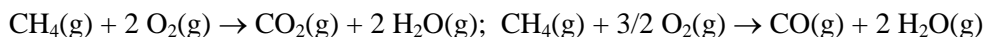
The balanced equation requires 2 mol  $\text{O}_2$  for every mol of  $\text{CH}_4$  that reacts. For three times as much oxygen, we would need 6 mol  $\text{O}_2$  per mol of  $\text{CH}_4$  reacted ( $n_{\text{O}_2} = 6n_{\text{CH}_4}$ ). Air is 21% mole percent  $\text{O}_2$ , so  $n_{\text{O}_2} = (0.21)n_{\text{air}}$ . Therefore, the moles of air we would need to deliver the excess  $\text{O}_2$  are:

$$n_{\text{O}_2} = (0.21)n_{\text{air}} = 6n_{\text{CH}_4}, \quad n_{\text{air}} = 29n_{\text{CH}_4}, \quad \frac{n_{\text{air}}}{n_{\text{CH}_4}} = 29$$

In 1 minute:

$$V_{\text{air}} = V_{\text{CH}_4} \times \frac{n_{\text{air}}}{n_{\text{CH}_4}} \times \frac{P_{\text{CH}_4}}{P_{\text{air}}} = 200. \text{ L} \times 29 \times \frac{1.50 \text{ atm}}{1.00 \text{ atm}} = 8.7 \times 10^3 \text{ L air/min}$$

- b. If  $x$  mol of  $\text{CH}_4$  were reacted, then  $6x$  mol  $\text{O}_2$  were added, producing  $(0.950)x$  mol  $\text{CO}_2$  and  $(0.050)x$  mol of  $\text{CO}$ . In addition,  $2x$  mol  $\text{H}_2\text{O}$  must be produced to balance the hydrogens.



Amount  $\text{O}_2$  reacted:

$$(0.950)x \text{ mol CO}_2 \times \frac{2 \text{ mol O}_2}{\text{mol CO}_2} = (1.90)x \text{ mol O}_2$$

$$(0.050)x \text{ mol CO} \times \frac{1.5 \text{ mol O}_2}{\text{mol CO}} = (0.075)x \text{ mol O}_2$$

$$\text{Amount of O}_2 \text{ left in reaction mixture} = (6.00)x - (1.90)x - (0.075)x = (4.03)x \text{ mol O}_2$$

$$\text{Amount of N}_2 = (6.00)x \text{ mol O}_2 \times \frac{79 \text{ mol N}_2}{21 \text{ mol O}_2} = (22.6)x \approx 23x \text{ mol N}_2$$

The reaction mixture contains:

$$(0.950)x \text{ mol CO}_2 + (0.050)x \text{ mol CO} + (4.03)x \text{ mol O}_2 + (2.00)x \text{ mol H}_2\text{O} \\ + 23x \text{ mol N}_2 = (30.)x \text{ mol of gas total}$$

$$\chi_{\text{CO}} = \frac{(0.050)x}{(30.)x} = 0.0017; \quad \chi_{\text{CO}_2} = \frac{(0.950)x}{(30.)x} = 0.032; \quad \chi_{\text{O}_2} = \frac{(4.03)x}{(30.)x} = 0.13$$

$$\chi_{\text{H}_2\text{O}} = \frac{(2.00)x}{(30.)x} = 0.067; \quad \chi_{\text{N}_2} = \frac{23x}{(30.)x} = 0.77$$

- c. The partial pressures are determined by  $P = \chi P_{\text{total}}$ . Because  $P_{\text{total}} = 1.00 \text{ atm}$ ,  $P_{\text{CO}} = 0.0017 \text{ atm}$ ,  $P_{\text{CO}_2} = 0.032 \text{ atm}$ ,  $P_{\text{O}_2} = 0.13 \text{ atm}$ ,  $P_{\text{H}_2\text{O}} = 0.067 \text{ atm}$ , and  $P_{\text{N}_2} = 0.77 \text{ atm}$ .

152.  $n_{\text{total}}$  = total number of moles of gas that have effused into the container:

$$n_{\text{total}} = \frac{PV}{RT} = \frac{(1.20 \times 10^{-6} \text{ atm}) \times 1.00 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 300. \text{ K}} = 4.87 \times 10^{-8} \text{ mol}$$

This amount has entered over a time span of 24 hours:

$$24 \text{ h} \times \frac{60 \text{ min}}{1 \text{ h}} \times \frac{60 \text{ s}}{1 \text{ min}} = 8.64 \times 10^4 \text{ s}$$

Thus:  $\frac{4.87 \times 10^{-8} \text{ mol}}{8.64 \times 10^4 \text{ s}} = 5.64 \times 10^{-13} \text{ mol/s}$  have entered the container.

$$\frac{5.64 \times 10^{-13} \text{ mol}}{\text{s}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} = 3.40 \times 10^{11} \text{ molecules/s}$$

The frequency of collisions of the gas with a given area is:

$$Z = A \left( \frac{N}{V} \right) \left( \frac{RT}{2\pi M} \right)^{1/2}; \quad Z_{\text{total}} = \frac{3.40 \times 10^{11} \text{ molecules}}{\text{s}} = Z_{\text{N}_2} + Z_{\text{O}_2}$$

$$\frac{n}{V} = \frac{P}{RT} = \frac{1.00 \text{ atm}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 300. \text{ K}} = 4.06 \times 10^{-2} \text{ mol/L}$$

$$\frac{N}{V} = \frac{4.06 \times 10^{-2} \text{ mol}}{\text{L}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} \times \frac{1000\text{L}}{\text{m}^3} = 2.44 \times 10^{25} \text{ molecules/m}^3$$

$$\text{For N}_2: \frac{N}{V} = (0.78)(2.44 \times 10^{25}) = 1.9 \times 10^{25} \text{ molecules/m}^3$$

$$\text{For O}_2: \frac{N}{V} = (0.22)(2.44 \times 10^{25}) = 5.4 \times 10^{24} \text{ molecules/m}^3$$

$$Z_{\text{total}} = 3.40 \times 10^{11} \text{ molecules/s} = Z_{\text{N}_2} + Z_{\text{O}_2}$$

$$3.40 \times 10^{11} = A \left[ 1.9 \times 10^{25} \left( \frac{8.3145 \times 300.}{2 \pi (28.0 \times 10^{-3})} \right)^{1/2} + 5.4 \times 10^{24} \left( \frac{8.3145 \times 300.}{2 \pi (32.0 \times 10^{-3})} \right)^{1/2} \right]$$

$$\frac{3.40 \times 10^{11} \text{ molecules}}{\text{s}} = A \left[ \frac{2.3 \times 10^{27} \text{ molecules}}{\text{m}^2 \text{ s}} + \frac{6.0 \times 10^{26} \text{ molecules}}{\text{m}^2 \text{ s}} \right]$$

$$A = \frac{3.40 \times 10^{11}}{2.9 \times 10^{27}} \text{ m}^2 = 1.2 \times 10^{-16} \text{ m}^2 = \pi r^2, \quad r = \left( \frac{1.2 \times 10^{-16} \text{ m}^2}{\pi} \right)^{1/2} = 6.2 \times 10^{-9} \text{ m} = 6.2 \text{ nm}$$

$$\text{Diameter of hole} = 2r = 2(6.2 \times 10^{-9} \text{ m}) = 1.2 \times 10^{-8} \text{ m} = 12 \text{ nm}$$

153. a. Average molar mass of air =  $0.790 \times 28.02 \text{ g/mol} + 0.210 \times 32.00 \text{ g/mol} = 28.9 \text{ g/mol}$ ;  
molar mass of helium =  $4.003 \text{ g/mol}$

A given volume of air at a given set of conditions has a larger density than helium at those conditions. We need to heat the air to a temperature greater than  $25^\circ\text{C}$  in order to lower the air density (by driving air out of the hot air balloon) until the density is the same as that for helium (at  $25^\circ\text{C}$  and  $1.00 \text{ atm}$ ).

- b. To provide the same lift as the helium balloon (assume  $V = 1.00 \text{ L}$ ), the mass of air in the hot-air balloon ( $V = 1.00 \text{ L}$ ) must be the same as that in the helium balloon. Let  $MM =$  molar mass:

$$P \cdot MM = dRT, \quad \text{mass} = \frac{MM \cdot PV}{RT}; \quad \text{solving: mass He} = 0.164 \text{ g}$$

$$\text{Mass air} = 0.164 \text{ g} = \frac{28.9 \text{ g/mol} \times 1.00 \text{ atm} \times 1.00 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times T}, \quad T = 2150 \text{ K} \quad (\text{a very high temperature})$$

**Marathon Problem**

154. We must determine the identities of element A and compound B in order to answer the questions. Use the first set of data to determine the identity of element A.

$$\text{Mass N}_2 = 659.452 \text{ g} - 658.572 \text{ g} = 0.880 \text{ g N}_2$$

$$0.880 \text{ g N}_2 \times \frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2} = 0.0314 \text{ mol N}_2$$

$$V = \frac{nRT}{P} = \frac{0.0314 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 288 \text{ K}}{790. \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}} = 0.714 \text{ L}$$

$$\text{Moles of A} = n = \frac{\left( 745 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} \right) \times 0.714 \text{ L}}{0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} \times (273 + 26) \text{ K}} = 0.0285 \text{ mol A}$$

$$\text{Mass of A} = 660.59 - 658.572 \text{ g} = 2.02 \text{ g A}$$

$$\text{Molar mass of A} = \frac{2.02 \text{ g A}}{0.0285 \text{ mol A}} = 70.9 \text{ g/mol}$$

The only element that is a gas at 26°C and 745 torr and has a molar mass close to 70.9 g/mol is chlorine = Cl<sub>2</sub> = element A.

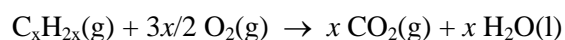
The remainder of the information is used to determine the formula of compound B. Assuming 100.00 g of B:

$$85.6 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 7.13 \text{ mol C}; \quad \frac{7.13}{7.13} = 1.00$$

$$14.4 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 14.3 \text{ mol H}; \quad \frac{14.13}{7.13} = 2.01$$

Empirical formula of B = CH<sub>2</sub>; molecular formula = C<sub>x</sub>H<sub>2x</sub> where x is a whole number.

The balanced combustion reaction of C<sub>x</sub>H<sub>2x</sub> with O<sub>2</sub> is:



To determine the formula of C<sub>x</sub>H<sub>2x</sub>, we need to determine the actual moles of all species present.

Mass of  $\text{CO}_2 + \text{H}_2\text{O}$  produced =  $846.7 \text{ g} - 765.3 \text{ g} = 81.4 \text{ g}$

Because  $\text{mol CO}_2 = \text{mol H}_2\text{O} = x$  (see balanced equation):

$$81.4 \text{ g} = x \text{ mol CO}_2 \times \frac{44.01 \text{ g CO}_2}{\text{mol CO}_2} + x \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g H}_2\text{O}}{\text{mol H}_2\text{O}}, \quad x = 1.31 \text{ mol}$$

$$\text{Mol O}_2 \text{ reacted} = 1.31 \text{ mol CO}_2 \times \frac{1.50 \text{ mol O}_2}{\text{mol CO}_2} = 1.97 \text{ mol O}_2$$

From the data, we can calculate the moles of excess  $\text{O}_2$  because only  $\text{O}_2(\text{g})$  remains after the combustion reaction has gone to completion.

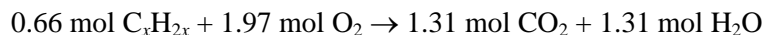
$$n_{\text{O}_2} = \frac{PV}{RT} = \frac{6.02 \text{ atm} \times 10.68 \text{ L}}{0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} \times (273 + 22) \text{ K}} = 2.66 \text{ mol excess O}_2$$

$$\text{Mol O}_2 \text{ present initially} = 1.97 \text{ mol} + 2.66 \text{ mol} = 4.63 \text{ mol O}_2$$

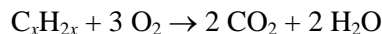
$$\text{Total moles gaseous reactants before reaction} = \frac{PV}{RT} = \frac{11.98 \text{ atm} \times 10.68 \text{ L}}{0.08206 \times 295 \text{ K}} = 5.29 \text{ mol}$$

$$\text{Mol C}_x\text{H}_{2x} = 5.29 \text{ mol total} - 4.63 \text{ mol O}_2 = 0.66 \text{ mol C}_x\text{H}_{2x}$$

Summarizing:

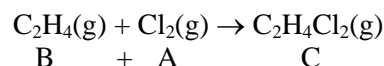


Dividing all quantities by 0.66 gives:



To balance the equation,  $\text{C}_x\text{H}_{2x}$  must be  $\text{C}_2\text{H}_4$  = compound B.

a. Now we can answer the questions. The reaction is:



$$\text{Mol Cl}_2 = n = \frac{PV}{RT} = \frac{1.00 \text{ atm} \times 10.0 \text{ L}}{0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 273 \text{ K}} = 0.446 \text{ mol Cl}_2$$

$$\text{Mol C}_2\text{H}_4 = n = \frac{PV}{RT} = \frac{1.00 \text{ atm} \times 8.60 \text{ L}}{0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 273 \text{ K}} = 0.384 \text{ mol C}_2\text{H}_4$$

Because a 1 : 1 mol ratio is required by the balanced reaction,  $\text{C}_2\text{H}_4$  is limiting.

$$\begin{aligned}\text{Mass C}_2\text{H}_4\text{Cl}_2 \text{ produced} &= 0.384 \text{ mol C}_2\text{H}_4 \times \frac{1 \text{ mol C}_2\text{H}_4\text{Cl}_2}{\text{mol C}_2\text{H}_4} \times \frac{98.95 \text{ g}}{\text{mol C}_2\text{H}_4\text{Cl}_2} \\ &= 38.0 \text{ g C}_2\text{H}_4\text{Cl}_2\end{aligned}$$

b. Excess mol  $\text{Cl}_2 = 0.446 \text{ mol Cl}_2 - 0.384 \text{ mol Cl}_2 \text{ reacted} = 0.062 \text{ mol Cl}_2$

$$P_{\text{total}} = \frac{n_{\text{total}}RT}{V}$$

$$n_{\text{total}} = 0.384 \text{ mol C}_2\text{H}_4\text{Cl}_2 \text{ produced} + 0.062 \text{ mol Cl}_2 \text{ excess} = 0.446 \text{ mol}$$

$$V = 10.0 \text{ L} + 8.60 \text{ L} = 18.6 \text{ L}$$

$$P_{\text{total}} = \frac{0.446 \text{ mol} \times 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 273 \text{ K}}{18.6 \text{ L}} = 0.537 \text{ atm}$$