

10 Gases

Visualizing Concepts

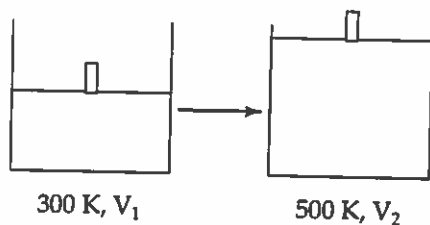
- 10.1 It would be much easier to drink from a straw on Mars. When a straw is placed in a glass of liquid, the liquid level in the straw equals the liquid level in the glass. The atmosphere exerts equal pressure inside and outside the straw. When we drink through a straw, we withdraw air, thereby reducing the pressure on the liquid inside. If only 0.007 atm is exerted on the liquid in the glass, a very small reduction in pressure inside the straw will cause the liquid to rise.

Another approach is to consider the gravitational force on Mars. Since the pull of gravity causes atmospheric pressure, the gravity on Mars must be much smaller than that on Earth. With a very small Martian gravity holding liquid in a glass, it would be very easy to raise the liquid through a straw.

- 10.2 (a) $V_1/T_1 = V_2/T_2$ (Charles' Law)

$$V_1/300 \text{ K} = V_2/500 \text{ K}$$

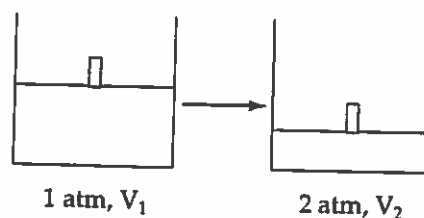
$$V_2 = 5/3 V_1$$



- (b) $P_1V_1 = P_2V_2$ (Boyle's Law)

$$1 \text{ atm} \times V_1 = 2 \text{ atm} \times V_2$$

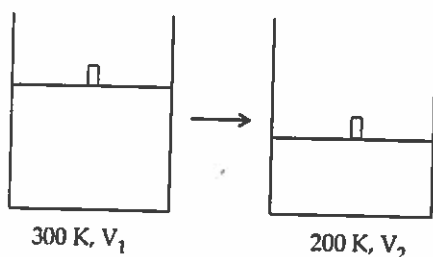
$$V_2 = 1/2 V_1$$



- (c) $V_1/T_1 = V_2/T_2$ (Charles' Law)

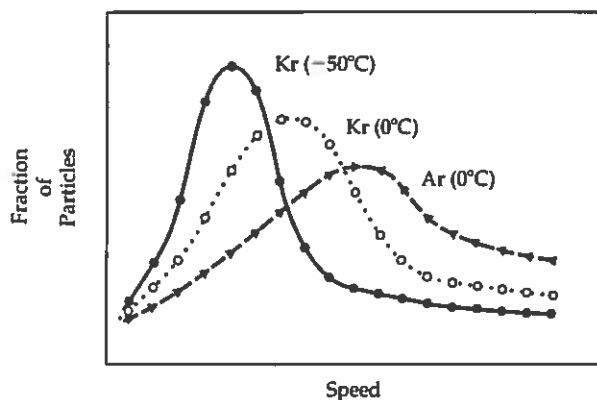
$$V_1/300 \text{ K} = V_2/200 \text{ K}$$

$$V_2 = 2/3 V_1$$

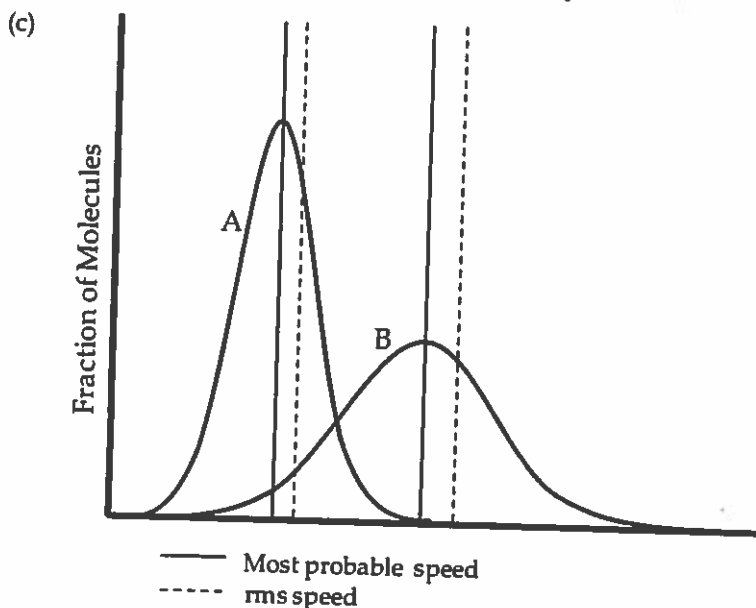


- 10.3 At constant temperature and volume, pressure depends on total number of particles (Charles' Law). In order to reduce the pressure by a factor of 2, the number of particles must be reduced by a factor of 2. At the lower pressure, the container would have half as many particles as at the higher pressure.
- 10.4 (a) At constant pressure and temperature, the container volume is directly proportional to the number of particles present (Avogadro's Law). As the reaction proceeds, 3 gas molecules are converted to 2 gas molecules, so the container volume decreases. If the reaction goes to completion, the final volume would be 2/3 of the initial volume.
- (b) At constant volume and temperature, pressure is directly proportional to the number of particles. Since the number of molecules decreases as the reaction proceeds, the pressure also decreases. At completion, the final pressure would be 2/3 the initial pressure.
- 10.5 $PV = nRT$ (ideal gas equation). In the ideal gas equation, R is a constant. Given constant V and n (fixed amount of ideal gas), P and T are directly proportional. If P is doubled, T is also doubled. That is, if P is doubled, T increases by a factor of two.
- 10.6 Over time, the gases will mix perfectly. Each bulb will contain 4 blue and 3 red atoms. The "blue" gas has the greater partial pressure after mixing, because it has the greater number of particles (at the same T and V as the "red" gas.)
- 10.7 (a) Partial pressure depends on the number of particles of each gas present. Red has the fewest particles, then yellow, then blue. $P_{\text{red}} < P_{\text{yellow}} < P_{\text{blue}}$
- (b) $P_{\text{gas}} = \chi_{\text{gas}} P_{\text{t}}$. Calculate the mole fraction, $\chi_{\text{gas}} = [\text{mol gas} / \text{total moles}]$ or $[\text{particles gas} / \text{total particles}]$. This is true because Avogadro's number is a counting number, and mole ratios are also particle ratios.
- $\chi_{\text{red}} = 2 \text{ red atoms} / 10 \text{ total atoms} = 0.2$; $P_{\text{red}} = 0.2(1.40 \text{ atm}) = 0.28 \text{ atm}$
- $\chi_{\text{yellow}} = 3 \text{ yellow atoms} / 10 \text{ total atoms} = 0.3$; $P_{\text{yellow}} = 0.3(1.40 \text{ atm}) = 0.42 \text{ atm}$
- $\chi_{\text{blue}} = 5 \text{ blue atoms} / 10 \text{ total atoms} = 0.5$; $P_{\text{blue}} = 0.5(1.40 \text{ atm}) = 0.70 \text{ atm}$
- Check. $(0.28 \text{ atm} + 0.42 \text{ atm} + 0.70 \text{ atm}) = 1.40 \text{ atm}$. The sum of the calculated partial pressures equals the given total pressure.

10.8



- 10.9 (a) At constant temperature, the root mean square (rms) speed (as well as the average speed) of a collection of gas particles is inversely related to molar mass; the lighter the particle, the faster it moves. Therefore, curve B represents He and curve A represents O_2 . Curve B has the higher rms speed and He is the lighter gas. Curve A has the lower rms speed and O_2 is the heavier gas.
- (b) For the same gas, average kinetic energy ($1/2 m u_{rms}^2$), and therefore root mean square speed (u_{rms}) is directly related to Kelvin temperature. Curve A is the lower temperature and curve B is the higher temperature.



- 10.10 (a) Total pressure is directly related to total number of particles (or total mol particles). $P(ii) < P(i) = P(iii)$
- (b) Partial pressure of He is directly related to number of He atoms (yellow) or mol He atoms. $P_{He}(iii) < P_{He}(ii) < P_{He}(i)$
- (c) Density is total mass of gas per unit volume. We can use the atomic or molar masses of He (4) and N_2 (28), as relative masses of the particles.

$$\text{mass}(i) = 5(4) + 2(28) = 76$$

$$\text{mass}(ii) = 3(4) + 1(28) = 40$$

$$\text{mass}(iii) = 2(4) + 5(28) = 148$$

Since the container volumes are equal, $d(ii) < d(i) < d(iii)$.

- (d) At the same temperature, all gases have the same average kinetic energy. The average kinetic energies of the particles in the three containers are equal.
- 10.11 The $NH_4Cl(s)$ ring will form at location A. The process described in this exercise is diffusion, rather than simple effusion. According to section 10.8, Graham's law approximates (but does not exactly describe) the diffusion rates of two gases under identical conditions. According to Graham's law, the ratio of rates is inversely related to the ratio of molar masses of the two gases. That is, the lighter gas moves faster than the

heavier gas. When introduced into the tube, NH_3 , $\text{MM} = 17$, moves faster and therefore farther than HCl , $\text{MM} = 36$. If NH_3 moves farther than HCl , the two gases meet and form $\text{NH}_4\text{Cl}(s)$ nearer the end where HCl was introduced; this is in the vicinity of location A.

10.12

CCl_4 , Cl_2 and H_2O would deviate most, while He , Ne and H_2 would deviate least. At high pressure, gas molecules have relatively little free space in which to move. Both intermolecular attractive forces and the nonzero volume of molecules are important. At low temperature, molecules have less kinetic energy, are less able to overcome attractive forces, and deviations due to intermolecular attractive forces dominate (see Figure 10.23). In Table 10.3, values of a show corrections for intermolecular attraction. The gases CCl_4 , Cl_2 and H_2O have the largest values of a , while He , Ne and H_2 have the smallest. The magnitude of a increases as the size and complexity of the molecule increases. CCl_4 and Cl_2 are relatively large gases; H_2O is polar and has hydrogen bonding. He , Ne and H_2 are small, monatomic or nonpolar diatomic. The magnitude of a is predictable from molecular structure.

Gas Characteristics; Pressure (sections 10.1 and 10.2)

- 10.13 In the gas phase, molecules are far apart and in constant motion. In the liquid phase, molecules maintain contact with each other while they move.
- A gas is much less dense than a liquid because most of the volume of a gas is empty space.
 - A gas is much more compressible because of the distance between molecules.
 - Gaseous molecules are so far apart that there is no barrier to mixing, regardless of the identity of the molecule. All mixtures of gases are homogeneous. Liquid molecules are touching. In order to mix, they must displace one another. Similar molecules displace each other and form homogeneous mixtures. Very dissimilar molecules form heterogeneous mixtures.
 - Gas molecules are far apart and constantly moving, so a gas expands to the shape and volume of its container. In a liquid sample, molecules are moving but in close contact with each other; there is very little empty space in a liquid. A liquid takes the shape of its container because the molecules are mobile, but maintains its own volume because the molecules are touching.
- 10.14
- Because gas molecules are far apart and in constant motion, the gas expands to fill the container. Attractive forces hold liquid molecules together and the volume of the liquid does not change.
 - H_2O and CCl_4 molecules are too dissimilar to displace each other and mix in the liquid state. All mixtures of gases are homogeneous. (See Solution 10.13 (c)).
 - Because gas molecules are far apart, the mass present in 1 mL of a gas is very small. The mass of a gas present in 1 L is on the same order of magnitude as the mass of a liquid present in 1 mL.
- 10.15 *Analyze.* Given: mass, area. Find: pressure. *Plan.* $P = F/A = m \times a/A$; use this relationship, paying attention to units. *Solve.*

$$(a) \quad 1 \text{ Pa} = \frac{1 \text{ N}}{\text{m}^2} = \frac{1 \text{ kg} \cdot \text{m}}{\text{s}^2} \times \frac{1}{\text{m}^2} = \frac{1 \text{ kg}}{\text{m} \cdot \text{s}^2} \quad \text{Change mass to kg and area to m}^2.$$

$$P = \frac{m \times a}{A} = \frac{130 \text{ lb}}{0.50 \text{ in}^2} \times \frac{9.81 \text{ m}}{1 \text{ s}^2} \times \frac{0.454 \text{ kg}}{1 \text{ lb}} \times \frac{39.4^2 \text{ in}^2}{1 \text{ m}^2} = 1.798 \times 10^6 \frac{\text{kg}}{\text{m} \cdot \text{s}^2} \\ = 1.8 \times 10^6 \text{ Pa} = 1.8 \times 10^3 \text{ kPa}$$

Check. $[1.30 \times 10 \times 0.5 \times (40)^2 / 0.5] \approx (130 \times 16,000) \approx 2.0 \times 10^6 \text{ Pa} \approx 2.0 \times 10^3 \text{ kPa}$.
The units are correct.

$$(b) \quad 1 \text{ atm} = 101.325 \text{ kPa}$$

$$1.798 \times 10^3 \text{ kPa} \times \frac{1 \text{ atm}}{101.325 \text{ kPa}} = 1.774 \times 10^1 = 18 \text{ atm}$$

$$(c) \quad 1 \text{ atm} = 14.70 \text{ lb/in}^2$$

$$17.74 \text{ atm} \times \frac{14.70 \text{ lb/in}^2}{1 \text{ atm}} = 260.8 = 2.6 \times 10^2 \text{ lb/in}^2$$

$$10.16 \quad P = m \times a / A; \quad 1 \text{ Pa} = 1 \text{ kg/m} \cdot \text{s}^2; \quad A = 3.0 \text{ cm} \times 4.1 \text{ cm} \times 4 = 49.2 = 49 \text{ cm}^2$$

$$\frac{262 \text{ kg}}{49.2 \text{ cm}^2} \times \frac{9.81 \text{ m}}{\text{s}^2} \times \frac{(100)^2 \text{ cm}^2}{1 \text{ m}^2} = 5.224 \times 10^5 \frac{\text{kg}}{\text{m} \cdot \text{s}^2} = 5.2 \times 10^5 \text{ Pa}$$

10.17 Analyze. Given: 760 mm column of Hg, densities of Hg and H₂O. Find: height of a column of H₂O at same pressure.

Plan. We must develop a relationship between pressure, height of a column of liquid, and density of the liquid. Relationships that might prove useful: $P = F/A$; $F = m \times a$; $m = d \times V$ (density)(volume); $V = A \times \text{height}$ Solve.

$$P = \frac{F}{A} = \frac{m \times a}{A} = \frac{d \times V \times a}{A} = \frac{d \times A \times h \times a}{A} = d \times h \times a$$

$$(a) \quad P_{\text{Hg}} = P_{\text{H}_2\text{O}}; \quad \text{Using the relationship derived above: } (d \times h \times a)_{\text{H}_2\text{O}} = (d \times h \times a)_{\text{Hg}}$$

Since a , the acceleration due to gravity, is equal in both liquids,

$$(d \times h)_{\text{H}_2\text{O}} = (d \times h)_{\text{Hg}}$$

$$1.00 \text{ g/mL} \times h_{\text{H}_2\text{O}} = 13.6 \text{ g/mL} \times 760 \text{ mm}$$

$$h_{\text{H}_2\text{O}} = \frac{13.6 \text{ g/mL} \times 760 \text{ mm}}{1.00 \text{ g/mL}} = 1.034 \times 10^4 = 1.03 \times 10^4 \text{ mm} = 10.3 \text{ m}$$

$$(b) \quad \text{Pressure due to H}_2\text{O:}$$

$$1 \text{ atm} = 1.034 \times 10^4 \text{ mm H}_2\text{O} \text{ (from part (a))}$$

$$39 \text{ ft H}_2\text{O} \times \frac{12 \text{ in}}{1 \text{ ft}} \times \frac{2.54 \text{ cm}}{1 \text{ in}} \times \frac{10 \text{ mm}}{1 \text{ cm}} \times \frac{1 \text{ atm}}{1.034 \times 10^4 \text{ mm}} = 1.150 = 1.2 \text{ atm}$$

$$P_{\text{total}} = P_{\text{atm}} + P_{\text{H}_2\text{O}} = 0.97 \text{ atm} + 1.150 \text{ atm} = 2.120 = 2.1 \text{ atm}$$

10.18 Using the relationship derived in Solution 10.17 for two liquids under the influence of gravity, $(d \times h)_{\text{lid}} = (d \times h)_{\text{Hg}}$. At 749 torr, the height of an Hg barometer is 749 mm.

$$\frac{1.20 \text{ g}}{1 \text{ mL}} \times h_{\text{lid}} = \frac{13.6 \text{ g}}{1 \text{ mL}} \times 760 \text{ mm}; \quad h_{\text{lid}} = \frac{13.6 \text{ g/mL} \times 749 \text{ mm}}{1.20 \text{ g/mL}} = 8.49 \times 10^3 \text{ mm} = 8.49 \text{ m}$$

- 10.19 (a) The tube can have any cross-sectional area. (The height of the Hg column in a barometer is independent of the cross-sectional area. See the expression for pressure derived in Solution 10.17.)
- (b) At equilibrium, the force of gravity per unit area acting on the mercury column is not equal to the force of gravity per unit area acting on the atmosphere. ($F = ma$; the acceleration due to gravity is equal for the two substances, but the mass of Hg for a given cross-sectional area is different than the mass of air for this same area.)
- (c) The column of mercury is held up by the pressure of the atmosphere applied to the exterior pool of mercury.
- (d) If you took the mercury barometer with you on a trip from the beach to high mountains, the height of the mercury column would decrease with elevation. (Atmospheric pressure decreases as elevation increases.)

10.20 The mercury would fill the tube completely; there would be no vacuum at the closed end. This is because atmospheric pressure will support a mercury column higher than 70 cm, while our tube is only 50 cm. No mercury flows from the tube into the dish and no vacuum forms at the top of the tube.

10.21 *Analyze/Plan.* Follow the logic in Sample Exercise 10.1. *Solve.*

$$(a) \quad 265 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.349 \text{ atm}$$

$$(b) \quad 265 \text{ torr} \times \frac{1 \text{ mm Hg}}{1 \text{ torr}} = 265 \text{ mm Hg}$$

$$(c) \quad 265 \text{ torr} \times \frac{1.01325 \times 10^5 \text{ Pa}}{760 \text{ torr}} = 3.53 \times 10^4 \text{ Pa}$$

$$(d) \quad 265 \text{ torr} \times \frac{1.01325 \times 10^5 \text{ Pa}}{760 \text{ torr}} \times \frac{1 \text{ bar}}{1 \times 10^5 \text{ Pa}} = 0.353 \text{ bar}$$

$$(e) \quad 265 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} \times \frac{14.70 \text{ psi}}{1 \text{ atm}} = 5.13 \text{ psi}$$

$$10.22 (a) \quad 0.912 \text{ atm} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 693 \text{ torr}$$

$$(b) \quad 0.685 \text{ bar} \times \frac{1 \times 10^5 \text{ Pa}}{1 \text{ bar}} \times \frac{1 \text{ kPa}}{1 \times 10^3 \text{ Pa}} = 68.5 \text{ kPa}$$

$$(c) \quad 655 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 0.862 \text{ atm}$$

$$(d) \quad 1.323 \times 10^5 \text{ Pa} \times \frac{1 \text{ atm}}{1.01325 \times 10^5 \text{ Pa}} = 1.3057 = 1.306 \text{ atm}$$

$$(e) \quad 2.50 \text{ atm} \times \frac{14.70 \text{ psi}}{1 \text{ atm}} = 36.75 = 36.8 \text{ psi}$$

10.23 *Analyze/Plan.* Follow the logic in Sample Exercise 10.1. *Solve.*

$$(a) \quad 30.45 \text{ in Hg} \times \frac{25.4 \text{ mm}}{1 \text{ in}} \times \frac{1 \text{ torr}}{1 \text{ mm Hg}} = 773.4 \text{ torr}$$

[The result has 4 sig figs because 25.4 mm/in is considered to be an exact number.]

(b) $30.45 \text{ in Hg} = 773.4 \text{ torr}; 773.4 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.018 \text{ atm}$

- (c) The pressure in Chicago is greater than standard atmospheric pressure, 760 torr or 1 atm, so it makes sense to classify this weather system as a "high pressure system."

10.24 $882 \text{ mbar} = 0.882 \text{ bar}$

(a) $0.882 \text{ bar} \times \frac{1 \times 10^5 \text{ Pa}}{1 \text{ bar}} \times \frac{1 \text{ atm}}{101,325 \text{ Pa}} = 0.8705 = 0.871 \text{ atm}$

(b) $0.882 \text{ bar} = 0.8705 \text{ atm} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 661.55 = 662 \text{ torr}$

(c) $0.882 \text{ bar} = 661.55 \text{ torr} \times \frac{1 \text{ mm Hg}}{1 \text{ torr}} \times \frac{1 \text{ cm Hg}}{10 \text{ mm Hg}} \times \frac{1 \text{ in Hg}}{2.54 \text{ cm Hg}} = 26.045 = 26.0 \text{ in Hg}$

10.25 *Analyze/Plan.* Follow the logic in Sample Exercise 10.2. *Solve.*

- (i) The Hg level is lower in the open end than the closed end, so the gas pressure is less than atmospheric pressure.

$$P_{\text{gas}} = 0.995 \text{ atm} - \left(52 \text{ cm} \times \frac{1 \text{ atm}}{76.0 \text{ cm}} \right) = 0.31 \text{ atm}$$

- (ii) The Hg level is higher in the open end, so the gas pressure is greater than atmospheric pressure.

$$P_{\text{gas}} = 0.995 \text{ atm} + \left(67 \text{ cm Hg} \times \frac{1 \text{ atm}}{76.0 \text{ cm Hg}} \right) = 1.8766 = 1.88 \text{ atm}$$

- (iii) This is a closed-end manometer, so $P_{\text{gas}} = h$.

$$P_{\text{gas}} = 10.3 \text{ cm} \times \frac{1 \text{ atm}}{76.0 \text{ cm}} = 0.136 \text{ atm}$$

- 10.26 (a) The atmosphere is exerting 15.4 cm = 154 mm Hg (torr) more pressure than the gas.

$$P_{\text{gas}} = P_{\text{atm}} - 15.4 \text{ torr} = \left(0.985 \text{ atm} \times \frac{760 \text{ torr}}{1 \text{ atm}} \right) - 15.4 \text{ torr} = 733 \text{ torr}$$

- (b) The gas is exerting 8.7 mm Hg (torr) more pressure than the atmosphere.

$$P_{\text{gas}} = P_{\text{atm}} + 12.3 \text{ torr} = \left(0.99 \text{ atm} \times \frac{760 \text{ torr}}{1 \text{ atm}} \right) + 12.3 \text{ torr} = 764.7 \text{ torr} = 7.6 \times 10^2 \text{ torr}$$

(Atmospheric pressure of 0.99 atm determines that the result has 2 sig figs.)

The Gas Laws (section 10.3)

10.27 *Analyze/Plan.* Given certain changes in gas conditions, predict the effect on other conditions. Consider the gas law relationships in section 10.3. *Solve.*

- (a) P and V are inversely proportional at constant T. If the volume decreases by a factor of 4, the pressure increases by a factor of 4.
- (b) P and T are directly proportional at constant V. If T decreases by a factor of 2, P also decreases by a factor of 2.

- (c) P and n are directly proportional at constant V and T . If n decreases by a factor of 4, P also decreases by a factor of 4.

10.28

Analyze. Given: initial P , V , T . Find: final values of P , V , T for certain changes of condition. *Plan.* Select the appropriate gas law relationships from Section 10.3; solve for final conditions, paying attention to units. *Solve.*

- (a) $P_1V_1 = P_2V_2$; the proportionality holds true for any pressure or volume units.

$$P_1 = 752 \text{ torr}, V_1 = 5.12 \text{ L}, P_2 = 1.88 \text{ atm}$$

$$V_2 = \frac{P_1V_1}{P_2} = \frac{752 \text{ torr} \times 5.12 \text{ L}}{1.88 \text{ atm}} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 2.69 \text{ L}$$

Check. As pressure increases, volume should decrease; our result agrees with this.

- (b) $V_1/T_1 = V_2/T_2$; T must be in Kelvins for the relationship to be true.

$$V_1 = 5.12 \text{ L}, T_1 = 21^\circ\text{C} = 294 \text{ K}, T_2 = 175^\circ\text{C} = 448 \text{ K}$$

$$V_2 = \frac{V_1T_2}{T_1} = \frac{5.12 \text{ L} \times 448 \text{ K}}{294 \text{ K}} = 7.80 \text{ L}$$

Check. As temperature increases, volume should increase; our result is consistent with this.

10.29

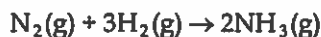
- (a) Avogadro's hypothesis states that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules. Since molecules react in the ratios of small whole numbers, it follows that the volumes of reacting gases (at the same temperature and pressure) are in the ratios of small whole numbers.

- (b) Since the two gases are at the same temperature and pressure, the ratio of the numbers of atoms is the same as the ratio of volumes. There are 1.5 times as many Xe atoms as Ne atoms.

- (c) Yes. By definition, one mole of an ideal gas (or any other substance) contains Avogadro's number of particles. At a given temperature and pressure, equal numbers of particles occupy the same volume, so one mole of an ideal gas will always occupy the same volume at the given temperature and pressure.

10.30

According to Avogadro's hypothesis, the mole ratios in the chemical equation will be volume ratios for the gases if they are at the same temperature and pressure.



The volumes of H_2 and N_2 are in a stoichiometric $\frac{3.6 \text{ L}}{1.2 \text{ L}}$ or $\frac{3 \text{ vol H}_2}{1 \text{ vol N}_2}$ ratio, so either can be used to determine the volume of $\text{NH}_3(\text{g})$ produced.

$$1.2 \text{ L N}_2 \times \frac{2 \text{ mol NH}_3}{1 \text{ mol N}_2} = 2.4 \text{ L NH}_3(\text{g}) \text{ produced.}$$

The Ideal-Gas Equation (section 10.4)

(In *Solutions to Exercises*, the symbol for molar mass is MM.)

- 10.31 (a) An ideal gas exhibits pressure, volume, and temperature relationships which are described by the equation $PV = nRT$. (An ideal gas obeys the ideal-gas equation.)

- (b) $V = \text{constant}/P$; Boyle's Law
 $V = \text{constant} \times T$; Charles Law
 $V = \text{constant} \times n$; Avogadro's Law

$$\text{Collecting all the equalities, } V = \frac{\text{constant} \times T \times n}{P}$$

If we call the constant R , and multiply both sides by P ,
 $PV = RTn$ or $PV = nRT$

- (c) $PV = nRT$; P in atmospheres, V in liters, n in moles, T in kelvins

$$(d) \quad R = \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{1.01325 \times 10^5 \text{ Pa}}{1 \text{ atm}} \times \frac{1 \text{ bar}}{1 \times 10^5 \text{ Pa}} = \frac{0.08315 \text{ L} \cdot \text{bar}}{\text{mol} \cdot \text{K}}$$

- 10.32 (a) STP stands for standard temperature, 0°C (or 273 K), and standard pressure, 1 atm .

$$(b) \quad V = \frac{nRT}{P}; V = 1 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{273 \text{ K}}{1 \text{ atm}}$$

$V = 22.4 \text{ L}$ for 1 mole of gas at STP

- (c) $25^\circ\text{C} + 273 = 298 \text{ K}$

$$V = \frac{nRT}{P}; V = 1 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{298 \text{ K}}{1 \text{ atm}}$$

$V = 24.5 \text{ L}$ for 1 mol of gas at 1 atm and 25°C

- 10.33 *Analyze/Plan.* $PV = nRT$. At constant volume and temperature, P is directly proportional to n .

Solve. For samples with equal masses of gas, the gas with $MM = 30$ will have twice as many moles of particles and twice the pressure. Thus, flask A contains the gas with $MM = 30$ and flask B contains the gas with $MM = 60$.

- 10.34 $n = g/MM$; $PV = nRT = gRT/MM$; $MM = gRT/PV$.

$$2\text{-L flask: } MM = 4.8 \text{ RT}/2.0(X) = 2.4 \text{ RT}/X$$

$$3\text{-L flask: } MM = 0.36 \text{ RT}/3.0 (0.1 X) = 1.2 \text{ RT}/X$$

The molar masses of the two gases are not equal. The gas in the 2-L flask has a molar mass that is twice as large as the gas in the 3-L flask.

- 10.35 *Analyze/Plan.* Follow the strategy for calculations involving many variables given in section 10.4. *Solve.*

$$T = \frac{PV}{nR} = 2.00 \text{ atm} \times \frac{1.00 \text{ L}}{0.500 \text{ mol}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} = 48.7 \text{ K}$$

$$K = 27^\circ\text{C} + 273 = 300 \text{ K}$$

$$n = \frac{PV}{RT} = 0.300 \text{ atm} \times \frac{0.250 \text{ L}}{300 \text{ K}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} = 3.05 \times 10^{-3} \text{ mol}$$

$$650 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.85526 = 0.855 \text{ atm}$$

$$V = \frac{nRT}{P} = 0.333 \text{ mol} \times \frac{350 \text{ K}}{0.85526 \text{ atm}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} = 11.2 \text{ L}$$

$$585 \text{ mL} = 0.585 \text{ L}$$

$$P = \frac{nRT}{V} = 0.250 \text{ mol} \times \frac{295 \text{ K}}{0.585 \text{ L}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} = 10.3 \text{ atm}$$

P	V	N	T
2.00 atm	1.00 L	0.500 mol	48.7 K
0.300 atm	0.250 L	3.05×10^{-3} mol	27°C
650 torr	11.2 L	0.333 mol	350 K
10.3 atm	585 mL	0.250 mol	295 K

10.36 *Analyze/Plan.* Follow the strategy for calculations involving many variables given in section 10.4. *Solve.*

(a) $n = 1.50 \text{ mol}$, $P = 1.25 \text{ atm}$, $T = -6^\circ\text{C} = 267 \text{ K}$

$$V = \frac{nRT}{P} = 1.50 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{267 \text{ K}}{1.25 \text{ atm}} = 26.3 \text{ L}$$

(b) $n = 3.33 \times 10^{-3} \text{ mol}$, $V = 478 \text{ mL} = 0.478 \text{ L}$

$$P = 750 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.9868 = 0.987 \text{ atm}$$

$$T = \frac{PV}{nR} = 0.9868 \text{ atm} \times \frac{0.478 \text{ L}}{3.33 \times 10^{-3} \text{ mol}} \times \frac{1 \text{ mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} = 1726 = 1.73 \times 10^3 \text{ K}$$

(c) $n = 0.00245 \text{ mol}$, $V = 413 \text{ mL} = 0.413 \text{ L}$, $T = 138^\circ\text{C} = 411 \text{ K}$

$$P = \frac{nRT}{V} = 0.00245 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{411 \text{ K}}{0.413 \text{ L}} = 0.200 \text{ atm}$$

(d) $V = 126.5 \text{ L}$, $T = 54^\circ\text{C} = 327 \text{ K}$,

$$P = 11.25 \text{ kPa} \times \frac{1 \text{ atm}}{101.325 \text{ kPa}} = 0.11103 = 0.1110 \text{ atm}$$

$$n = \frac{PV}{RT} = 0.11103 \text{ atm} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{126.5 \text{ L}}{327 \text{ K}} = 0.523 \text{ mol}$$

10.37 *Analyze/Plan.* Follow the strategy for calculations involving many variables. *Solve.*

$$n = g/\text{MM}; PV = nRT; PV = gRT/\text{MM}; g = \text{MM} \times PV/RT$$

$P = 1.0 \text{ atm}$, $T = 23^\circ\text{C} = 296 \text{ K}$, $V = 1.75 \times 10^5 \text{ ft}^3$. Change ft^3 to L, then calculate grams (or kg).

$$1.75 \times 10^5 \text{ ft}^3 \times \frac{(12)^3 \text{ in}^3}{\text{ft}^3} \times \frac{(2.54)^3 \text{ cm}^3}{\text{in}^3} \times \frac{1 \text{ L}}{1 \times 10^3 \text{ cm}^3} = 4.9554 \times 10^6 = 4.96 \times 10^6 \text{ L}$$

$$g = \frac{4.003 \text{ g He}}{1 \text{ mol He}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{1.0 \text{ atm} \times 4.955 \times 10^6 \text{ L}}{296 \text{ K}} = 8.2 \times 10^5 \text{ g} = 820 \text{ kg He}$$

10.38 Find the volume of the tube in cm^3 ; $1 \text{ cm}^3 = 1 \text{ mL}$.

$$r = d/2 = 2.5 \text{ cm}/2 = 1.25 = 1.3 \text{ cm}; h = 5.5 \text{ m} = 5.5 \times 10^2 \text{ cm}$$

$$V = \pi r^2 h = 3.14159 \times (1.25 \text{ cm})^2 \times (5.5 \times 10^2 \text{ cm}) = 2.700 \times 10^3 \text{ cm}^3 = 2.7 \text{ L}$$

$$PV = \frac{g}{MM} RT; g = \frac{MM \times PV}{RT}; P = 1.78 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 2.342 \times 10^{-3} = 2.34 \times 10^{-3} \text{ atm}$$

$$g = \frac{20.18 \text{ g Ne}}{1 \text{ mol Ne}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{2.342 \times 10^{-3} \text{ atm} \times 2.700 \text{ L}}{308 \text{ K}} = 5.049 \times 10^{-3} = 5.0 \times 10^{-3} \text{ g Ne}$$

10.39 *Analyze/Plan.* Follow the strategy for calculations involving many variables. *Solve.*

$$(a) V = 2.25 \text{ L}; T = 273 + 37^\circ\text{C} = 310 \text{ K}; P = 735 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.96710 = 0.967 \text{ atm}$$

$$PV = nRT, n = PV/RT, \text{ number of molecules (\#)} = n \times 6.022 \times 10^{23}$$

$$\# = \frac{0.9671 \text{ atm} \times 2.25 \text{ L}}{310 \text{ K}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} = 5.15 \times 10^{22} \text{ molecules}$$

$$(b) V = 5.0 \times 10^3 \text{ L}; T = 273 + 0^\circ\text{C} = 273 \text{ K}; P = 1.00 \text{ atm}; MM = 28.98 \text{ g/mol}$$

$$PV = \frac{g}{MM} RT; g = \frac{MM \times PV}{RT}$$

$$g = \frac{28.98 \text{ g air}}{1 \text{ mol air}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{1.00 \text{ atm} \times 5.0 \times 10^3 \text{ L}}{273 \text{ K}}$$

$$= 6,468 \text{ g} = 6.5 \times 10^3 \text{ g air} = 6.5 \text{ kg air}$$

$$10.40 (a) P_{\text{O}_3} = 3.0 \times 10^{-3} \text{ atm}; T = 250 \text{ K}; V = 1 \text{ L (exact)}$$

$$\# \text{ of O}_3 \text{ molecules} = \frac{PV}{RT} \times 6.022 \times 10^{23}$$

$$\# = \frac{3.0 \times 10^{-3} \text{ atm} \times 1 \text{ L}}{250 \text{ K}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}}$$

$$= 8.8 \times 10^{19} \text{ O}_3 \text{ molecules}$$

$$(b) \# \text{ of CO}_2 \text{ molecules} = \frac{PV}{RT} \times 6.022 \times 10^{23} \times 0.0004$$

$$\# = \frac{1.0 \text{ atm} \times 2.0 \text{ L}}{300 \text{ K}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} \times 0.0004$$

$$= 1.957 \times 10^{19} = 2 \times 10^{19} \text{ CO}_2 \text{ molecules}$$

10.41 *Analyze/Plan.* Follow the strategy for calculations involving many variables. *Solve.*

$$(a) P = \frac{nRT}{V}; n = 0.29 \text{ kg O}_2 \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 9.0625 = 9.1 \text{ mol}; V = 2.3 \text{ L};$$

$$T = 273 + 9^\circ\text{C} = 282 \text{ K}$$

$$P = \frac{9.0625 \text{ mol}}{2.3 \text{ L}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 282 \text{ K} = 91 \text{ atm}$$

$$(b) V = \frac{nRT}{P}; = \frac{9.0625 \text{ mol}}{0.95 \text{ atm}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 299 \text{ K} = 2.3 \times 10^2 \text{ L}$$

$$10.42 (a) V = 0.250 \text{ L}, T = 23^\circ\text{C} = 296 \text{ K}, n = 2.30 \text{ g C}_3\text{H}_8 \times \frac{1 \text{ mol C}_3\text{H}_8}{44.1 \text{ g C}_3\text{H}_8} = 0.052154 = 0.0522 \text{ mol}$$

$$P = \frac{nRT}{V} = 0.052154 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{296 \text{ K}}{0.250 \text{ L}} = 5.07 \text{ atm}$$

(b) STP = 1.00 atm, 273 K

$$V = \frac{nRT}{P} = 0.052154 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{273 \text{ K}}{1.00 \text{ atm}} = 1.1684 = 1.17 \text{ L}$$

(c) $^{\circ}\text{C} = 5/9 (^{\circ}\text{F} - 32^{\circ})$; $\text{K} = ^{\circ}\text{C} + 273.15 = 5/9 (130^{\circ}\text{F} - 32^{\circ}) + 273.15 = 327.59 = 328 \text{ K}$

$$P = \frac{nRT}{V} = 0.052154 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{327.59 \text{ K}}{0.250 \text{ L}} = 5.608 = 5.61 \text{ atm}$$

10.43 *Analyze/Plan.* Follow the strategy for calculations involving many variables. *Solve.*

$$V = 8.70 \text{ L}, T = 24^{\circ}\text{C} = 297 \text{ K}, P = 895 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.1776 = 1.18 \text{ atm}$$

$$(a) \quad g = \frac{MM \times PV}{RT}; g = \frac{70.91 \text{ g Cl}_2}{1 \text{ mol Cl}_2} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{1.1776 \text{ atm}}{297 \text{ K}} \times 8.70 \text{ L} \\ = 29.8 \text{ g Cl}_2$$

$$(b) \quad V_2 = \frac{P_1 V_1 T_2}{T_1 P_2} = \frac{895 \text{ torr} \times 8.70 \text{ L} \times 273 \text{ K}}{297 \text{ K} \times 760 \text{ torr}} = 9.42 \text{ L}$$

$$(c) \quad T_2 = \frac{P_2 V_2 T_1}{P_1 V_1} = \frac{876 \text{ torr} \times 15.00 \text{ L} \times 297 \text{ K}}{895 \text{ torr} \times 8.70 \text{ L}} = 501 \text{ K}$$

$$(d) \quad P_2 = \frac{P_1 V_1 T_2}{V_2 T_1} = \frac{895 \text{ torr} \times 8.70 \text{ L} \times 331 \text{ K}}{5.00 \text{ L} \times 297 \text{ K}} = 1.73 \times 10^3 \text{ torr} = 2.28 \text{ atm}$$

10.44 $T = 23^{\circ}\text{C} = 296 \text{ K}, P = 16,500 \text{ kPa} \times \frac{1 \text{ atm}}{101.325 \text{ kPa}} = 162.84 = 163 \text{ atm}$

$$V = 55.0 \text{ gal} \times \frac{3.7854 \text{ L}}{1 \text{ gal}} = 208.20 = 208 \text{ L}$$

$$(a) \quad g = \frac{MM \times PV}{RT}; g = \frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{162.84 \text{ atm}}{296 \text{ K}} \times 208.20 \text{ L} \\ = 4.4665 \times 10^4 \text{ g O}_2 = 44.7 \text{ kg O}_2$$

$$(b) \quad V_2 = \frac{P_1 V_1 T_2}{T_1 P_2} = \frac{16,500 \text{ kPa} \times 208.20 \text{ L} \times 273 \text{ K}}{296 \text{ K} \times 101.325 \text{ kPa}} = 3.13 \times 10^4 \text{ L}$$

$$(c) \quad T_2 = \frac{P_2 T_1}{P_1} = \frac{150.0 \text{ atm} \times 296 \text{ K}}{16,500 \text{ kPa}} \times \frac{101.325 \text{ kPa}}{1 \text{ atm}} = 272.7 = 273 \text{ K}$$

$$(d) \quad P_2 = \frac{P_1 V_1 T_2}{V_2 T_1} = \frac{16,500 \text{ kPa} \times 208.20 \text{ L} \times 297 \text{ K}}{55.0 \text{ L} \times 296 \text{ K}} = 62,671 = 6.27 \times 10^4 \text{ kPa}$$

10.45 *Analyze.* Given: mass of cockroach, rate of O_2 consumption, temperature, percent O_2 in air, volume of air. Find: mol O_2 consumed per hour; mol O_2 in 1 quart of air; mol O_2 consumed in 48 hr.

(a) *Plan/Solve.* V of O_2 consumed = rate of consumption \times mass \times time. $n = PV/RT$.

$$5.2 \text{ g} \times 1 \text{ hr} \times \frac{0.8 \text{ mL O}_2}{1 \text{ g} \cdot \text{hr}} = 4.16 = 4 \text{ mL O}_2 \text{ consumed}$$

$$(b) \quad n = \frac{PV}{RT} = 1 \text{ atm} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{0.00416 \text{ L}}{297 \text{ K}} = 1.71 \times 10^{-4} = 2 \times 10^{-4} \text{ mol O}_2$$

- (c) *Plan/Solve.* qt air \rightarrow L air \rightarrow L O₂ available. mol O₂ available = PV/RT.
mol O₂/hr (from part (a)) \rightarrow total mol O₂ consumed. Compare O₂ available and O₂ consumed.

$$1 \text{ qt air} \times \frac{0.946 \text{ L}}{1 \text{ qt}} \times 0.21 \text{ O}_2 \text{ in air} = 0.199 \text{ L O}_2 \text{ available}$$

$$n = 1 \text{ atm} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{0.199 \text{ L}}{297 \text{ K}} = 8.17 \times 10^{-3} = 8 \times 10^{-3} \text{ mol O}_2 \text{ available}$$

$$\text{roach uses} \frac{1.71 \times 10^{-4} \text{ mol}}{1 \text{ hr}} \times 48 \text{ hr} = 8.21 \times 10^{-3} = 8 \times 10^{-3} \text{ mol O}_2 \text{ consumed}$$

Not only does the roach use 20% of the available O₂, it needs all the O₂ in the jar.

- 10.46 Change mass to kg; 1 hr = 60 min; pay attention to units.

$$(a) \quad 185 \text{ lb} \times \frac{1 \text{ kg}}{2.2046 \text{ lb}} \times \frac{47.5 \text{ mL O}_2}{\text{kg} \cdot \text{min}} \times 60 \text{ min} = 2.39 \times 10^5 \text{ mL}$$

$$(b) \quad 165 \text{ lb} \times \frac{1 \text{ kg}}{2.2046 \text{ lb}} \times \frac{65.0 \text{ mL O}_2}{\text{kg} \cdot \text{min}} \times 60 \text{ min} = 2.92 \times 10^5 \text{ mL}$$

- 10.47 (a) *Analyze.* Given: 119 tons Hg, 1 atm, 298 K. Find: vol Hg(g). *Plan.* Change tons to grams; use $V = gRT/MM \times P$ to calculate volume. *Solve.*

$$119 \text{ tons} \times \frac{2000 \text{ lb}}{1 \text{ ton}} \times \frac{453.59 \text{ g}}{1 \text{ lb}} = 1.07954 \times 10^8 = 1.08 \times 10^8 \text{ g Hg}$$

$$V = \frac{1.07954 \times 10^8 \text{ g Hg} \times 1 \text{ mol Hg}}{200.59 \text{ g Hg}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{298 \text{ K}}{1 \text{ atm}}$$

$$= 1.3161 \times 10^7 = 1.32 \times 10^7 \text{ L}$$

- (b) *Analyze.* Given: total vol atmospheric gases (air) = $51 \times 10^{12} \text{ m}^3$; 245 ppb Hg(g) by volume. Find: mol Hg(g) in atmospheres. *Plan.* Change m^3 to L; use definition of ppb to get L Hg(g); use $n = PV/RT$ to calculate mol Hg. Assume STP. *Solve.*

$$51 \times 10^{12} \text{ m}^3 \times \frac{10^3 \text{ dm}^3}{1 \text{ m}^3} \times \frac{1 \text{ L}}{1 \text{ dm}^3} = 5.1 \times 10^{16} \text{ L air}$$

$$5.1 \times 10^{16} \text{ L air} \times \frac{245 \text{ L Hg(g)}}{1 \times 10^9 \text{ L air}} = 1.2495 \times 10^{10} = 1.2 \times 10^{10} \text{ L Hg(g)}$$

$$n = \frac{1 \text{ atm} \times 1.2495 \times 10^{10} \text{ L Hg(g)}}{273 \text{ K}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} = 5.5775 \times 10^8 = 5.6 \times 10^8 \text{ mol Hg(g)}$$

Check. Note that calculating amount of Hg from volume Hg(g) depends on the the assumed temperature of the atmosphere.

- 10.48 mass = $1800 \times 10^{-9} \text{ g} = 1.8 \times 10^{-6} \text{ g}$; $V = 1 \text{ m}^3 = 1 \times 10^3 \text{ L}$; $T = 273 + 10^\circ \text{C} = 283 \text{ K}$

$$(a) \quad P = \frac{gRT}{MM \times V}; P = \frac{1.8 \times 10^{-6} \text{ g Hg} \times 1 \text{ mol Hg}}{200.6 \text{ g Hg}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{283 \text{ K}}{1 \times 10^3 \text{ L}}$$

$$= 2.1 \times 10^{-10} \text{ atm}$$

$$(b) \frac{1.8 \times 10^{-6} \text{ g Hg}}{1 \text{ m}^3} \times \frac{1 \text{ mol Hg}}{200.6 \text{ g Hg}} \times \frac{6.022 \times 10^{23} \text{ Hg atoms}}{1 \text{ mol Hg}} = 5.4 \times 10^{15} \text{ Hg atoms/m}^3$$

$$(c) 1600 \text{ km}^3 \times \frac{1000^3 \text{ m}^3}{1 \text{ km}^3} \times \frac{1.8 \times 10^{-6} \text{ g Hg}}{1 \text{ m}^3} = 2.9 \times 10^6 \text{ g Hg/day}$$

Further Applications of the Ideal-Gas Equation (section 10.5)

10.49 $\text{Cl}_2(\text{g})$ is the most dense at 1.00 atm and 298 K. Gas density is directly proportional to molar mass and pressure, and inversely proportional to temperature (Equation [10.10]). For gas samples at the same conditions, molar mass determines density. Of the three gases listed, Cl_2 has the largest molar mass.

10.50 $\text{CO}_2 < \text{SO}_2 < \text{HBr}$. For gases at the same conditions, density is directly proportional to molar mass. The order of increasing molar mass is the order of increasing density. CO_2 , 44 g/mol < SO_2 , 64 g/mol < HBr , 81 g/mol.

10.51 (c) Because the helium atoms are of lower mass than the average air molecule, the helium gas is less dense than air. The balloon thus weighs less than the air displaced by its volume.

10.52 (b) Xe atoms have a higher mass than N_2 molecules. Because both gases at STP have the same number of molecules per unit volume, the Xe gas must be denser.

10.53 *Analyze/Plan.* Conditions (P, V, T) and amounts of gases are given. Rearrange the relationship $PV \times \text{MM} = gRT$ to obtain the desired of quantity, paying attention (as always!) to units. *Solve.*

$$(a) d = \frac{\text{MM} \times P}{RT}; \text{MM} = 46.0 \text{ g/mol}; P = 0.970 \text{ atm}, T = 35^\circ\text{C} = 308 \text{ K}$$

$$d = \frac{46.0 \text{ g NO}_2}{1 \text{ mol}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{0.970 \text{ atm}}{308 \text{ K}} = 1.77 \text{ g/L}$$

$$(b) \text{MM} = \frac{gRT}{PV} = \frac{2.50 \text{ g}}{0.875 \text{ L}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{308 \text{ K}}{685 \text{ torr}} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 80.1 \text{ g/mol}$$

$$10.54 (a) d = \frac{\text{MM} \times P}{RT}; \text{MM} = 146.1 \text{ g/mol}, T = 21^\circ\text{C} = 294 \text{ K}, P = 707 \text{ torr}$$

$$d = \frac{146.1 \text{ g}}{1 \text{ mol}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{707 \text{ torr}}{294 \text{ K}} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 5.63 \text{ g/L}$$

$$(b) \text{MM} = \frac{dRT}{P} = \frac{7.135 \text{ g}}{1 \text{ L}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{285 \text{ K}}{743 \text{ torr}} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 171 \text{ g/mol}$$

10.55 *Analyze/Plan.* Given: mass, conditions (P, V, T) of unknown gas. Find: molar mass. $\text{MM} = gRT/PV$. *Solve.*

$$\text{MM} = \frac{gRT}{PV} = \frac{1.012 \text{ g}}{0.354 \text{ L}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{372 \text{ K}}{742 \text{ torr}} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 89.4 \text{ g/mol}$$

$$10.56 \text{MM} = \frac{gRT}{PV} = \frac{0.846 \text{ g}}{0.354 \text{ L}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{373 \text{ K}}{752 \text{ torr}} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 73.9 \text{ g/mol}$$

10.57 *Analyze/Plan.* Follow the logic in Sample Exercise 10.9. *Solve.*

$$\text{mol O}_2 = \frac{PV}{RT} = 3.5 \times 10^{-6} \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{0.452 \text{ L}}{300 \text{ K}} = 8.456 \times 10^{-11}$$

$$= 8.5 \times 10^{-11} \text{ mol O}_2$$

$$8.456 \times 10^{-11} \text{ mol O}_2 \times \frac{2 \text{ mol Mg}}{1 \text{ mol O}_2} \times \frac{24.3 \text{ g Mg}}{1 \text{ mol Mg}} = 4.1 \times 10^{-9} \text{ g Mg (4.1 ng Mg)}$$

10.58 $n_{\text{H}_2} = \frac{P_{\text{H}_2} V}{RT}$; $P = 825 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.0855 = 1.09 \text{ atm}$; $T = 273 + 21^\circ\text{C} = 294 \text{ K}$

$$n_{\text{H}_2} = 1.0855 \text{ atm} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{145 \text{ L}}{294 \text{ K}} = 6.5242 = 6.52 \text{ mol H}_2$$

$$6.5242 \text{ mol H}_2 \times \frac{1 \text{ mol CaH}_2}{2 \text{ mol H}_2} \times \frac{42.10 \text{ g CaH}_2}{1 \text{ mol CaH}_2} = 137.34 = 137 \text{ g CaH}_2$$

10.59 (a) *Analyze/Plan.* g glucose \rightarrow mol glucose \rightarrow mol CO₂ \rightarrow V CO₂ *Solve.*

$$24.5 \text{ g} \times \frac{1 \text{ mol glucose}}{180.1 \text{ g}} \times \frac{6 \text{ mol CO}_2}{1 \text{ mol glucose}} = 0.8162 = 0.816 \text{ mol CO}_2$$

$$V = \frac{nRT}{P} = 0.8162 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{310 \text{ K}}{0.970 \text{ atm}} = 21.4 \text{ L CO}_2$$

(b) *Analyze/Plan.* g glucose \rightarrow mol glucose \rightarrow mol O₂ \rightarrow V O₂ *Solve.*

$$50.0 \text{ g} \times \frac{1 \text{ mol glucose}}{180.1 \text{ g}} \times \frac{6 \text{ mol O}_2}{1 \text{ mol glucose}} = 1.6657 = 1.67 \text{ mol O}_2$$

$$V = \frac{nRT}{P} = 1.6657 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{298 \text{ K}}{1 \text{ atm}} = 40.7 \text{ L O}_2$$

10.60 Follow the logic in Sample Exercise 10.9. The H₂(g) will be used in a balloon, which operates at atmospheric pressure. Since atmospheric pressure is not explicitly given, assume 1 atm (infinite sig figs).

$$n = \frac{PV}{RT} = 1 \text{ atm} \times \frac{3.1150 \times 10^4 \text{ L}}{295 \text{ K}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} = 1.28678 \times 10^3 = 1.29 \times 10^3 \text{ mol H}_2$$

From the balanced equation, 1 mol of Fe produces 1 mol of H₂, so 1.29 \times 10³ mol Fe are required.

$$1.28678 \times 10^3 \text{ mol Fe} \times \frac{55.845 \text{ g Fe}}{\text{mol Fe}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 71.86 = 71.9 \text{ kg Fe}$$

10.61 *Analyze/Plan.* The gas sample is a mixture of H₂(g) and H₂O(g). Find the partial pressure of H₂(g) and then the moles of H₂(g) and Zn(s). *Solve.*

$$P_1 = 738 \text{ torr} = P_{\text{H}_2} + P_{\text{H}_2\text{O}}$$

From Appendix B, the vapor pressure of H₂O at 24°C = 22.38 torr

$$P_{\text{H}_2} = (738 \text{ torr} - 22.38 \text{ torr}) \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.9416 = 0.942 \text{ atm}$$

$$n_{\text{H}_2} = \frac{P_{\text{H}_2} V}{RT} = 0.9416 \text{ atm} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{0.159 \text{ L}}{297 \text{ K}} = 0.006143 = 0.00614 \text{ mol H}_2$$

$$0.006143 \text{ mol H}_2 \times \frac{1 \text{ mol Zn}}{1 \text{ mol H}_2} \times \frac{65.39 \text{ g Zn}}{1 \text{ mol Zn}} = 0.402 \text{ g Zn}$$

10.62 The gas sample is a mixture of $C_2H_2(g)$ and $H_2O(g)$. Find the partial pressure of C_2H_2 , then moles CaC_2 and C_2H_2 .

$$P_t = 753 \text{ torr} = P_{C_2H_2} + P_{H_2O} \quad P_{H_2O} \text{ at } 23^\circ\text{C} = 21.07 \text{ torr}$$

$$P_{C_2H_2} = (753 \text{ torr} - 21.07 \text{ torr}) \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.96307 = 0.963 \text{ atm}$$

$$1.524 \text{ g } CaC_2 \times \frac{1 \text{ mol } CaC_2}{64.10 \text{ g}} \times \frac{1 \text{ mol } C_2H_2}{1 \text{ mol } CaC_2} = 0.023775 = 0.02378 \text{ mol } C_2H_2$$

$$V = 0.023775 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{296 \text{ K}}{0.96307 \text{ atm}} = 0.600 \text{ L } C_2H_2$$

Partial Pressures (section 10.6)

10.63 (a) When the stopcock is opened, the volume occupied by $N_2(g)$ increases from 2.0 L to 5.0 L. At constant T, $P_1V_1 = P_2V_2$. $1.0 \text{ atm} \times 2.0 \text{ L} = P_2 \times 5.0 \text{ L}$; $P_2 = 0.40 \text{ atm}$

(b) When the gases mix, the volume of $O_2(g)$ increases from 3.0 L to 5.0 L. At constant T, $P_1V_1 = P_2V_2$. $2.0 \text{ atm} \times 3.0 \text{ L} = P_2 \times 5.0 \text{ L}$; $P_2 = 1.2 \text{ atm}$

(c) $P_t = P_{N_2} + P_{O_2} = 0.40 \text{ atm} + 1.2 \text{ atm} = 1.6 \text{ atm}$

10.64 (a) The partial pressure of gas A is not affected by the addition of gas C. The partial pressure of A depends only on moles of A, volume of container, and conditions; none of these factors changes when gas C is added.

(b) The total pressure in the vessel increases when gas C is added, because the total number of moles of gas increases.

(c) The mole fraction of gas B decreases when gas C is added. The moles of gas B stay the same, but the total moles increase, so the mole fraction of B (n_B/n_t) decreases.

10.65 Analyze. Given: amount, V, T of three gases. Find: P of each gas, total P.

Plan. $P = nRT/V$; $P_t = P_1 + P_2 + P_3 + \dots$ Solve.

(a) $P_{He} = \frac{nRT}{V} = 0.765 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{298 \text{ K}}{10.00 \text{ L}} = 1.871 = 1.87 \text{ atm}$

$$P_{Ne} = \frac{nRT}{V} = 0.330 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{298 \text{ K}}{10.00 \text{ L}} = 0.8070 = 0.807 \text{ atm}$$

$$P_{Ar} = \frac{nRT}{V} = 0.110 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{298 \text{ K}}{10.00 \text{ L}} = 0.2690 = 0.269 \text{ atm}$$

(b) $P_t = 1.871 \text{ atm} + 0.8070 \text{ atm} + 0.2690 \text{ atm} = 2.9470 = 2.95 \text{ atm}$

10.66 Given mass, V and T of O_2 and He, find the partial pressure of each gas. Sum to find the total pressure in the tank.

$$V = 10.0 \text{ L}; T = 19^\circ\text{C}; 19 + 273 = 292 \text{ K}$$

$$n_{O_2} = 51.2 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{31.999 \text{ g } O_2} = 1.600 = 1.60 \text{ mol } O_2$$

$$n_{\text{He}} = 32.6 \text{ g He} \times \frac{1 \text{ mol He}}{4.0026 \text{ g He}} = 8.1447 = 8.14 \text{ mol He}$$

$$P_{\text{O}_2} = 1.600 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{292 \text{ K}}{10.0 \text{ L}} = 3.8338 = 3.84 \text{ atm}$$

$$P_{\text{He}} = 8.1447 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{292 \text{ K}}{10.0 \text{ L}} = 19.5159 = 19.5 \text{ atm}$$

$$P_t = 3.8338 + 19.5159 = 23.3497 = 23.3 \text{ atm}$$

- 10.67 *Analyze.* Given 390 ppm CO₂ in the atmosphere; 390 L CO₂ in 10⁶ total L air. Find: mole fraction CO₂ in the atmosphere. *Plan.* Avogadro's law deals with the relationship between volume and moles of a gas.

Solve. Avogadro's law states that volume of a gas at constant temperature and pressure is directly proportional to moles of the gas. Using volume fraction to express concentration assumes that the 390 L CO₂ and 10⁶ total L air are at the same temperature and pressure. That is, 390 L is the volume that the number of moles of CO₂ present in 10⁶ L air would occupy at atmospheric temperature and pressure. The mole fraction of CO₂ in the atmosphere is then just the volume fraction from the concentration by volume.

$$\chi_{\text{CO}_2} = \frac{390 \text{ L CO}_2}{10^6 \text{ L air}} = 0.00039$$

10.68 $\chi_{\text{Xe}} = 4/100 = 0.04$; $\chi_{\text{Ne}} = \chi_{\text{He}} = (1 - 0.04)/2 = 0.48$

$$V_1 = 0.900 \text{ mm} \times 0.300 \text{ mm} \times 10.0 \text{ mm} \times \frac{1 \text{ cm}^3}{10^3 \text{ mm}^3} \times \frac{1 \text{ L}}{1000 \text{ cm}^3} = 2.70 \times 10^{-6} \text{ L}$$

$$P_1 = 500 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.657895 = 0.658 \text{ atm}$$

$$n_t = \frac{PV}{RT} = 0.657895 \text{ atm} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{2.70 \times 10^{-6} \text{ L}}{298 \text{ K}} \times \frac{6.022 \times 10^{23} \text{ atoms}}{\text{mol}}$$

$$= 4.3743 \times 10^{16} = 4.37 \times 10^{16} \text{ total atoms}$$

$$\text{Xe atoms} = \chi_{\text{Xe}} \times \text{total atoms} = 0.04(4.3743 \times 10^{16}) = 1.75 \times 10^{15} = 2 \times 10^{15} \text{ Xe atoms}$$

$$\text{Ne atoms} = \text{He atoms} = 0.48(4.3743 \times 10^{16}) = 2.10 \times 10^{16} = 2.1 \times 10^{16} \text{ Ne and He atoms}$$

Assumptions: In order to calculate total moles of gas and total atoms, we assumed a reasonable room temperature. Since '4% Xe' was not defined, we conveniently assumed mole percent. The 1:1 relationship of Ne to He is assumed to be by volume and not by mass.

- 10.69 *Analyze.* Given: mass CO₂ at V, T; pressure of air at same V, T. Find: partial pressure of CO₂ at these conditions, total pressure of gases at V, T.

Plan. $\text{g CO}_2 \rightarrow \text{mol CO}_2 \rightarrow P_{\text{CO}_2}$ (via $P = nRT/V$); $P_t = P_{\text{CO}_2} + P_{\text{air}}$ *Solve.*

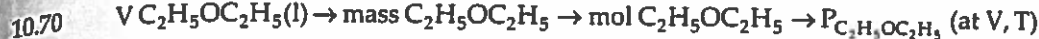
$$5.50 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} = 0.12497 = 0.125 \text{ mol CO}_2; T = 273 + 24^\circ \text{C} = 297 \text{ K}$$

$$P_{\text{CO}_2} = 0.12497 \text{ mol} \times \frac{297 \text{ K}}{10.0 \text{ L}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} = 0.30458 = 0.305 \text{ atm}$$

$$P_{\text{air}} = 705 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.92763 = 0.928 \text{ atm}$$

$$P_t = P_{\text{CO}_2} + P_{\text{air}} = 0.30458 + 0.92763 = 1.23221 = 1.232 \text{ atm}$$

(Result has 3 decimal places and 4 sig figs.)



$$P_t = P_{\text{N}_2} + P_{\text{O}_2} + P_{\text{C}_2\text{H}_5\text{OC}_2\text{H}_5}; T = 273.15 + 35.0^\circ\text{C} = 308.15 = 308.2 \text{ K}$$

$$\begin{aligned} \text{(a)} \quad 5.00 \text{ mL C}_2\text{H}_5\text{OC}_2\text{H}_5 &\times \frac{0.7134 \text{ g C}_2\text{H}_5\text{OC}_2\text{H}_5}{\text{mL}} \times \frac{1 \text{ mol C}_2\text{H}_5\text{OC}_2\text{H}_5}{74.12 \text{ g C}_2\text{H}_5\text{OC}_2\text{H}_5} \\ &= 0.048125 = 0.0481 \text{ mol C}_2\text{H}_5\text{OC}_2\text{H}_5 \end{aligned}$$

$$P = \frac{nRT}{V} = 0.048125 \text{ mol} \times \frac{308.15}{6.00 \text{ L}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} = 0.20282 = 0.203 \text{ atm}$$

$$\text{(b)} \quad P_t = P_{\text{N}_2} + P_{\text{O}_2} + P_{\text{C}_2\text{H}_5\text{OC}_2\text{H}_5} = 0.751 \text{ atm} + 0.208 \text{ atm} + 0.203 \text{ atm} = 1.162 \text{ atm}$$

10.71 *Analyze/Plan.* The partial pressure of each component is equal to the mole fraction of that gas times the total pressure of the mixture. Find the mole fraction of each component and then its partial pressure. *Solve.*

$$n_t = 0.75 \text{ mol N}_2 + 0.30 \text{ mol O}_2 + 0.15 \text{ mol CO}_2 = 1.20 \text{ mol}$$

$$\chi_{\text{N}_2} = \frac{0.75}{1.20} = 0.625 = 0.63; P_{\text{N}_2} = 0.625 \times 2.15 \text{ atm} = 1.344 = 1.3 \text{ atm}$$

$$\chi_{\text{O}_2} = \frac{0.30}{1.20} = 0.250 = 0.25; P_{\text{O}_2} = 0.250 \times 2.15 \text{ atm} = 0.5375 = 0.54 \text{ atm}$$

$$\chi_{\text{CO}_2} = \frac{0.15}{1.20} = 0.125 = 0.13; P_{\text{CO}_2} = 0.125 \times 2.15 \text{ atm} = 0.2688 = 0.27 \text{ atm}$$

Check. $P_T = 1.344 + 0.5375 + 0.2688 = 2.1503 = 2.15 \text{ atm}$. The sum of the partial pressures agrees with the stated total pressure.

10.72 $n_{\text{N}_2} = 10.25 \text{ g N}_2 \times \frac{1 \text{ mol}}{28.02 \text{ g}} = 0.3658 \text{ mol}; n_{\text{H}_2} = 1.83 \text{ g H}_2 \times \frac{1 \text{ mol}}{2.016 \text{ g}} = 0.9077 = 0.908 \text{ mol}$

$$n_{\text{NH}_3} = 7.95 \text{ g NH}_3 \times \frac{1 \text{ mol}}{17.03 \text{ g}} = 0.4668 \text{ mol} = 0.467 \text{ mol}$$

$$n_t = 0.3658 + 0.9077 + 0.4668 = 1.7403 = 1.740 \text{ mol}$$

$$P_{\text{N}_2} = \frac{n_{\text{N}_2}}{n_t} \times P_t = \frac{0.3658}{1.7403} \times 1.85 \text{ atm} = 0.389 \text{ atm}$$

$$P_{\text{H}_2} = \frac{0.9077}{1.7403} \times 1.85 \text{ atm} = 0.965 \text{ atm}; P_{\text{NH}_3} = \frac{0.4668}{1.7403} \times 1.85 \text{ atm} = 0.496 \text{ atm}$$

10.73 *Analyze/Plan.* Mole fraction = pressure fraction. Find the desired mole fraction of O₂ and change to mole percent. *Solve.*

$$\chi_{\text{O}_2} = \frac{P_{\text{O}_2}}{P_t} = \frac{0.21 \text{ atm}}{8.38 \text{ atm}} = 0.025; \text{mole \%} = 0.025 \times 100 = 2.5\%$$

$$10.74 \quad (a) \quad n_{\text{O}_2} = 15.08 \text{ g O}_2 \times \frac{1 \text{ mol}}{31.999 \text{ g}} = 0.4713 \text{ mol}; n_{\text{N}_2} = 8.17 \text{ g N}_2 \times \frac{1 \text{ mol}}{28.02 \text{ g}} = 0.292 \text{ mol}$$

$$n_{\text{H}_2} = 2.64 \text{ g H}_2 \times \frac{1 \text{ mol}}{2.016 \text{ g}} = 1.31 \text{ mol}; n_t = 0.4713 + 0.292 + 1.31 = 2.07 \text{ mol}$$

$$\chi_{\text{O}_2} = \frac{n_{\text{O}_2}}{n_t} = \frac{0.4713}{2.07} = 0.228; \chi_{\text{N}_2} = \frac{n_{\text{N}_2}}{n_t} = \frac{0.292}{2.07} = 0.141$$

$$\chi_{\text{H}_2} = \frac{1.31}{2.07} = 0.633$$

$$(b) \quad P_{\text{O}_2} = n \times \frac{RT}{V}; P_{\text{O}_2} = 0.4713 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{288.15 \text{ K}}{15.50 \text{ L}} = 0.7190 \text{ atm}$$

$$P_{\text{N}_2} = 0.292 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{288.15 \text{ K}}{15.50 \text{ L}} = 0.445 \text{ atm}$$

$$P_{\text{H}_2} = 1.31 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{288.15 \text{ K}}{15.50 \text{ L}} = 2.00 \text{ atm}$$

10.75 *Analyze/Plan.* $\text{N}_2(\text{g})$ and $\text{O}_2(\text{g})$ undergo changes of conditions and are mixed. Calculate the new pressure of each gas and add them to obtain the total pressure of the mixture.

$$P_2 = P_1 V_1 T_2 / V_2 T_1; P_T = P_{\text{N}_2} + P_{\text{O}_2}. \quad \text{Solve.}$$

$$P_{\text{N}_2} = \frac{P_1 V_1 T_2}{V_2 T_1} = \frac{5.25 \text{ atm} \times 1.00 \text{ L} \times 293 \text{ K}}{12.5 \text{ L} \times 299 \text{ K}} = 0.41157 = 0.412 \text{ atm}$$

$$P_{\text{O}_2} = \frac{P_1 V_1 T_2}{V_2 T_1} = \frac{5.25 \text{ atm} \times 5.00 \text{ L} \times 293 \text{ K}}{12.5 \text{ L} \times 299 \text{ K}} = 2.05786 = 2.06 \text{ atm}$$

$$P_t = 0.41157 \text{ atm} + 2.05786 \text{ atm} = 2.46943 = 2.47 \text{ atm}$$

10.76 Calculate the pressure of the gas in the second vessel directly from mass and conditions using the ideal-gas equation.

$$(a) \quad P_{\text{SO}_2} = \frac{gRT}{MV} = \frac{3.00 \text{ g SO}_2}{64.07 \text{ g SO}_2/\text{mol}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{299 \text{ K}}{10.0 \text{ L}} = 0.11489 = 0.115 \text{ atm}$$

$$(b) \quad P_{\text{N}_2} = \frac{gRT}{MV} = \frac{2.35 \text{ g N}_2}{28.01 \text{ g N}_2/\text{mol}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{299 \text{ K}}{10.0 \text{ L}} = 0.20585 = 0.206 \text{ atm}$$

$$(c) \quad P_t = P_{\text{SO}_2} + P_{\text{N}_2} = 0.11489 \text{ atm} + 0.20585 \text{ atm} = 0.321 \text{ atm}$$

Kinetic-Molecular Theory of Gases; Effusion and Diffusion (sections 10.7 and 10.8)

- 10.77 (a) Increase in temperature at constant volume, decrease in volume, increase in pressure
 (b) Decrease in temperature
 (c) Increase in volume, decrease in pressure
 (d) Increase in temperature

- 10.78 (a) False. The average kinetic energy per molecule in a collection of gas molecules is the same for all gases at the same temperature.
- (b) True.
- (c) False. The molecules in a gas sample at a given temperature exhibit a distribution of kinetic energies.
- (d) True.
- (e) False. Gas molecules at the same temperature exhibit a distribution of speeds.
- 10.79 The fact that gases are readily compressible supports the assumption that most of the volume of a gas sample is empty space.
- 10.80 Newton's model provides no explanation of the effect of a change in temperature on the pressure of a gas at constant volume or on the volume of a gas at constant pressure. On the other hand, the assumption that the average kinetic energy of gas molecules increases with increasing temperature explains Charles' Law, that an increase in temperature requires an increase in volume to maintain constant pressure.
- 10.81 Average speed is the numerical mean speed, the sum of the speeds of all particles divided by the total number of particles. The root mean square (rms) speed is the speed of a molecule with the same kinetic energy as the average kinetic energy of the sample. For a given gas sample at a fixed temperature, rms speed is larger (greater) than average speed, but the difference between the two is small.
- 10.82 The gas undergoes a chemical reaction which has fewer gas particles in products than in reactants. Mass is conserved when a chemical reaction occurs, so the mass of (flask + contents) remains constant. Pressure is directly proportional to number of particles, so pressure decreases as the number of gaseous particles decreases. One simple example of such a reaction is the dimerization of NO_2 : $2 \text{NO}_2(\text{g}) \rightarrow \text{N}_2\text{O}_4$.
- 10.83 *Analyze/Plan.* Apply the concepts of the Kinetic-Molecular Theory (KMT) to the situation where a gas is heated at constant volume. Determine how the quantities in (a)-(d) are affected by this change. *Solve.*
- (a) Average kinetic energy is proportional to temperature (K), so average kinetic energy of the molecules increases.
- (b) The average kinetic energy of a gas is $\frac{1}{2} m u_{\text{rms}}^2$. Molecular mass doesn't change as T increases; average kinetic energy increases so rms speed (u) increases. (Also, $u_{\text{rms}} = (3RT/MM)^{1/2}$, so u_{rms} is directly related to T.)
- (c) As T and thus rms molecular speed increase, molecular momentum (mu) increases and the strength of an average impact with the container wall increases.
- (d) As T and rms molecular speed increase, the molecules collide more frequently with the container walls, and the total number of collisions per second increases.
- 10.84 (a) They have the same number of molecules (equal volumes of gases at the same temperature and pressure contain equal numbers of molecules).
- (b) N_2 is more dense because it has the larger molar mass. Since the volumes of the samples and the number of molecules are equal, the gas with the larger molar mass will have the greater density.

- (c) The average kinetic energies are equal (statement 5, section 10.7).
 (d) CH_4 will effuse faster. The lighter the gas molecules, the faster they will effuse (Graham's Law).

- 10.85 (a) *Plan.* The larger the molar mass, the slower the average speed (at constant temperature).

Solve. In order of increasing speed (and decreasing molar mass):
 $\text{HBr} < \text{NF}_3 < \text{SO}_2 < \text{CO} < \text{Ne}$

- (b) *Plan.* Follow the logic of Sample Exercise 10.14. *Solve.*

$$u_{\text{rms}} = \sqrt{\frac{3RT}{\text{MM}}} = \left(\frac{3 \times 8.314 \text{ kg} \cdot \text{m}^2/\text{s}^2 \cdot \text{mol} \cdot \text{K} \times 298 \text{ K}}{71.0 \times 10^{-3} \text{ kg/mol}} \right)^{1/2} = 324 \text{ m/s}$$

- (c) *Plan.* Use Equation 10.23 to calculate the most probable speed, u_{mp} . MM of $\text{O}_3 = 48.0 \text{ g/mol}$; $T = 270 \text{ K}$. *Solve.*

$$u_{\text{mp}} = \sqrt{\frac{2RT}{\text{MM}}} = \left(\frac{2 \times 8.314 \text{ kg} \cdot \text{m}^2/\text{s}^2 \cdot \text{mol} \cdot \text{K} \times 270 \text{ K}}{48.0 \times 10^{-3} \text{ kg/mol}} \right)^{1/2} = 306 \text{ m/s}$$

- 10.86 (a) *Plan.* The greater the molecular (and molar) mass, the smaller the rms and average speeds of the molecules. Calculate the molar mass of each gas, and place them in decreasing order of mass and increasing order of rms and average speed. *Solve.*

$\text{CO} = 28 \text{ g/mol}$; $\text{SF}_6 = 146 \text{ g/mol}$; $\text{H}_2\text{S} = 34 \text{ g/mol}$; $\text{Cl}_2 = 71 \text{ g/mol}$;
 $\text{HBr} = 81 \text{ g/mol}$. In order of increasing speed (and decreasing molar mass):
 $\text{SF}_6 < \text{HBr} < \text{Cl}_2 < \text{H}_2\text{S} < \text{CO}$

- (b) *Plan.* Follow the logic of Sample Exercise 10.14. *Solve.*

$$\text{CO: } u_{\text{rms}} = \sqrt{\frac{3RT}{\text{MM}}} = \left(\frac{3 \times 8.314 \text{ kg} \cdot \text{m}^2/\text{s}^2 \cdot \text{mol} \cdot \text{K} \times 300 \text{ K}}{28.0 \times 10^{-3} \text{ kg/mol}} \right)^{1/2} = 5.17 \times 10^2 \text{ m/s}$$

$$\text{Cl}_2: u_{\text{rms}} = \left(\frac{3 \times 8.314 \text{ kg} \cdot \text{m}^2/\text{s}^2 \cdot \text{mol} \cdot \text{K} \times 300 \text{ K}}{70.9 \times 10^{-3} \text{ kg/mol}} \right)^{1/2} = 3.25 \times 10^2 \text{ m/s}$$

As expected, the lighter molecule moves at the greater speed.

- (c) *Plan.* From Equations [10.22] and [10.23], we see that the ratio of most probable speed to rms speed is $(2/3)^{1/2}$. Use this ratio and the results from part (b) to calculate most probable speeds. *Solve.*

$$\text{CO: } u_{\text{mp}} = (2/3)^{1/2}(5.17 \times 10^2 \text{ m/s}) = 422 \text{ m/s}$$

$$\text{Cl}_2: u_{\text{mp}} = (2/3)^{1/2}(3.25 \times 10^2 \text{ m/s}) = 265 \text{ m/s}$$

The lighter molecule, CO, has the greater most probable speed. Note that the most probable speed is less than the rms speed, as shown on Figure 10.17(b).

- 10.87 *Effusion* is the escape of gas molecules through a tiny hole. *Diffusion* is the distribution of a gas throughout space or throughout another substance. On a macroscopic scale,

effusion is like hundreds of students leaving an auditorium through one small door, while diffusion is analogous to these students making their way around the quad, which may or may not already contain other students.

- 10.88 Write each proportionality relationship as an equation, then combine them to obtain a formula for mean free path.

The operational symbols and units are: mean free path, λ , meters (m); temperature, T, kelvins (K); pressure, P, atmospheres (atm); diameter of a gas molecule, d, meters (m), constant, R_{mfp} .

$$\lambda = \text{constant} \times T; \lambda = \text{constant}/P; \lambda = \text{constant}/d^2$$

$$\text{Combining: } \lambda = \frac{R_{\text{mfp}} \times T}{P \times d^2}$$

The units of R_{mfp} are chosen and arranged so that they cancel the units of measurement, leaving an appropriate length unit for λ .

$$\text{With the units defined above, } R_{\text{mfp}} \text{ will have units of } \frac{\text{atm} \cdot \text{m}^3}{\text{K}}.$$

(Note that $1 \text{ m}^3 = 10^3 \text{ dm}^3 = 1000 \text{ L}$. Substituting, R_{mfp} would have units of $\frac{\text{atm} \cdot \text{L}}{\text{K}}$, with the factor of 1000 incorporated into the value of R_{mfp} .)

- 10.89 *Plan.* The heavier the molecule, the slower the rate of effusion. Thus, the order for increasing rate of effusion is in the order of decreasing mass. *Solve.*

$$\text{rate } ^2\text{H}^{37}\text{Cl} < \text{rate } ^1\text{H}^{37}\text{Cl} < \text{rate } ^2\text{H}^{35}\text{Cl} < \text{rate } ^1\text{H}^{35}\text{Cl}$$

$$10.90 \frac{\text{rate}^{235}\text{U}}{\text{rate}^{238}\text{U}} = \sqrt{\frac{238.05}{235.04}} = \sqrt{1.0128} = 1.0064$$

There is a slightly greater rate enhancement for $^{235}\text{U}(\text{g})$ atoms than $^{235}\text{UF}_6(\text{g})$ molecules (1.0043), because ^{235}U is a greater percentage (100%) of the mass of the diffusing particles than in $^{235}\text{UF}_6$ molecules. The masses of the isotopes were taken from *The Handbook of Chemistry and Physics*.

- 10.91 *Analyze.* Given: relative effusion rates of two gases at same temperature. Find: molecular formula of one of the gases. *Plan.* Use Graham's law to calculate the formula weight of arsenic (III) sulfide, and thus the molecular formula. *Solve.*

$$\frac{\text{rate (sulfide)}}{\text{rate (Ar)}} = \left[\frac{39.9}{\text{MM (sulfide)}} \right]^{1/2} = 0.28$$

$$\text{MM (sulfide)} = 39.9 / 0.28^2 = 509 \text{ g/mol (two significant figures)}$$

The empirical formula of arsenic(III) sulfide is As_2S_3 , which has a formula mass of 246.1. Twice this is 490 g/mol, close to the value estimated from the effusion experiment. Thus, the formula of the gas phase molecule is As_4S_6 .

- 10.92 The time required is proportional to the reciprocal of the effusion rate.

$$\frac{\text{rate (X)}}{\text{rate (O}_2)} = \frac{105 \text{ s}}{31 \text{ s}} = \left[\frac{32 \text{ g O}_2}{\text{MM}_x} \right]^{1/2}; \text{MM}_x = 32 \text{ g O}_2 \times \left[\frac{105}{31} \right]^2 = 370 \text{ g/mol (two sig figs)}$$

Nonideal-Gas Behavior (section 10.9)

- 10.93 (a) Nonideal gas behavior is observed at very high pressures and/or low temperatures.
- (b) The real volumes of gas molecules and attractive intermolecular forces between molecules cause gases to behave nonideally.
- (c) The ratio PV/RT is equal to the number of moles of particles in an ideal-gas sample; this number should be a constant for all pressure, volume, and temperature conditions. If the value of this ratio changes with increasing pressure, the gas sample is not behaving ideally. That is, the gas is not behaving according to the ideal-gas equation.

Negative deviations indicate fewer "effective" particles in the sample, a result of attractive forces among particles. Positive deviations indicate more "effective" particles, a result of the real volume occupied by the particles.

- 10.94 Ideal-gas behavior is most likely to occur at high temperature and low pressure, so the atmosphere on Mercury is more likely to obey the ideal-gas law. The higher temperature on Mercury means that the kinetic energies of the molecules will be larger relative to intermolecular attractive forces. Further, the gravitational attractive forces on Mercury are lower because the planet has a much smaller mass. This means that for the same column mass of gas (Figure 10.1), atmospheric pressure on Mercury will be lower.
- 10.95 *Plan.* The constants a and b are part of the correction terms in the van der Waals equation. The smaller the values of a and b , the smaller the corrections and the more ideal the gas. *Solve.*
- Ar ($a = 1.34$, $b = 0.0322$) will behave more like an ideal gas than CO_2 ($a = 3.59$, $b = 0.0427$) at high pressures.
- 10.96 The constant a is a measure of the strength of intermolecular attractions among gas molecules; b is a measure of molecular volume. Both increase with increasing molecular mass and structural complexity.
- 10.97 *Analyze/Plan.* Follow the logic in Sample Exercise 10.16. Use the ideal-gas equation to calculate pressure in (a), the van der Waals equation in (b). $n = 1.00$ mol, $V = 5.00$ L, $T = 25^\circ\text{C} = 298$ K; $a = 6.49$ L²-atm/mol², $b = 0.0562$ L/mol.
- (a)
$$P = \frac{nRT}{V} = 1.00 \text{ mol} \times \frac{298 \text{ K}}{5.00 \text{ L}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} = 4.89 \text{ atm}$$
- (b)
$$P = \frac{nRT}{V - nb} - \frac{n^2 a}{V^2};$$

$$P = \frac{(1.00 \text{ mol})(298 \text{ K})(0.08206 \text{ L} \cdot \text{atm}/\text{mol} \cdot \text{K})}{5.00 \text{ L} - (1.00 \text{ mol})(0.0562 \text{ L}/\text{mol})} - \frac{(1.00 \text{ mol})^2(6.49 \text{ L}^2 \cdot \text{atm}/\text{mol}^2)}{(5.00 \text{ L})^2}$$

$$P = 4.9463 \text{ atm} - 0.2596 \text{ atm} = 4.6868 = 4.69 \text{ atm}$$
- (c) From Sample Exercise 10.16, the difference at 22.41 L between the ideal and van der Waals results is $(1.00 - 0.990) = 0.010$ atm. At 5.00 L, the difference is $(4.89 - 4.69) = 0.20$ atm. The effects of both molecular attractions, the a correction, and molecular volume, the b correction, increase with decreasing volume. For the a correction, V^2 appears in the denominator, so the correction increases

exponentially as V decreases. For the b correction, nb is a larger portion of the total volume as V decreases. That is, 0.0562 L is 1.1% of 5.0 L, but only 0.25% of 22.41 L. Qualitatively, molecular attractions are more important as the amount of free space decreases and the number of molecular collisions increase. Molecular volume is a larger part of the total volume as the container volume decreases.

- 10.98 *Analyze.* Conditions and amount of $\text{CCl}_4(\text{g})$ are given. *Plan.* Use ideal-gas equation and van der Waals equation to calculate pressure of gas at these conditions. *Solve.*

$$(a) \quad P = 1.00 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{313 \text{ K}}{33.3 \text{ L}} = 0.771 \text{ atm}$$

$$(b) \quad P = \frac{nRT}{V - nb} - \frac{an^2}{V^2} = \frac{1.00 \times 0.08206 \times 313}{33.3 - (1.00 \times 0.1383)} - \frac{20.4(1.00)^2}{(33.3)^2} = 0.756 \text{ atm}$$

Check. The van der Waals result indicates that the real pressure will be less than the ideal pressure. That is, intermolecular forces reduce the effective number of particles and the real pressure. This is reasonable for 1 mole of gas at relatively low temperature and pressure.

- (c) According to Table 10.3, CCl_4 has larger a and b values. That is, CCl_4 experiences stronger intermolecular attractions and has a larger molecular volume than Cl_2 does. CCl_4 will deviate more from ideal behavior at these conditions than Cl_2 will.
- 10.99 *Analyze.* Given the b value of Xe, 0.0510 L/mol, calculate the radius of a Xe atom.

Plan. Use Avogadro's number to change L/mol to L/atom. Use the volume formula, $V = 4/3 \pi r^3$ and units conversion to obtain the radius in Å. 1 L = 1 dm³. *Solve.*

$$\frac{0.0510 \text{ L}}{1 \text{ mol Xe}} \times \frac{1 \text{ mol Xe}}{6.022 \times 10^{23} \text{ Xe atoms}} \times \frac{1 \text{ dm}^3}{1 \text{ L}} = 8.4689 \times 10^{-26} = 8.47 \times 10^{-26} \text{ dm}^3$$

$$V = 4/3 \pi r^3; \quad r^3 = 3V/4\pi; \quad r = (3V/4\pi)^{1/3}$$

$$r = \left(\frac{3 \times 8.4689 \times 10^{-26} \text{ dm}^3}{4 \times 3.14159} \right)^{1/3} = 2.7243 \times 10^{-9} = 2.72 \times 10^{-9} \text{ dm}$$

$$2.72 \times 10^{-9} \text{ dm} \times \frac{1 \text{ m}}{10 \text{ dm}} \times \frac{1 \text{ \AA}}{1 \times 10^{-10} \text{ m}} = 2.72 \text{ \AA}$$

The calculated value is the nonbonding radius. From Figure 7.6 in section 7.3, the bonding atomic radius of Xe is 1.30 Å. We expect the nonbonding radius of an atom to be larger than the bonding radius, but our calculated value is more than twice as large.

- 10.100 From section 7.3, the nonbonding or *van der Waals* radius is half of the shortest internuclear distance when two nonbonding atoms collide. So, radii calculated from the van der Waals equation are nonbonding radii. According to the kinetic molecular theory, ideal gas particles undergo perfectly elastic, billiard-ball collisions, in keeping with the definition of nonbonding radii.

Also, from the results of Exercise 10.99, the atomic radius calculated from the van der Waals b value is twice as large as the bonding atomic radius from Figure 7.6. Nonbonding radii are larger than bonding radii because no lasting penetration of electron clouds occurs during a nonbonding collision.

Additional Exercises

$$10.101 \quad P_1 V_1 = P_2 V_2; V_2 = P_1 V_1 / P_2$$

$$V_2 = \frac{3.0 \text{ atm} \times 1.0 \text{ mm}^3}{730 \text{ torr}} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 3.1 \text{ mm}^3$$

$$10.102 \quad PV = nRT, n = PV/RT. \text{ Since } RT \text{ is constant, } n \text{ is proportional to } PV.$$

$$\text{Total available } n = (15.0 \text{ L} \times 1.00 \times 10^2 \text{ atm}) - (15.0 \text{ L} \times 1.00 \text{ atm}) = 1485$$

$$\text{Each balloon holds } 2.00 \text{ L} \times 1.00 \text{ atm} = 2.00 \text{ L-atm} \qquad = 1.49 \times 10^3 \text{ L-atm}$$

$$\frac{1485 \text{ L-atm available}}{2.00 \text{ L-atm/balloon}} = 742.5 = 742 \text{ balloons}$$

(Only 742 balloons can be filled completely, with a bit of He left over.)

$$10.103 \quad P = \frac{nRT}{V}; n = 1.4 \times 10^{-5} \text{ mol}, V = 0.600 \text{ L}, T = 23^\circ\text{C} = 296 \text{ K}$$

$$P = 1.4 \times 10^{-5} \text{ mol} \times \frac{0.08206 \text{ L-atm}}{\text{mol-K}} \times \frac{296 \text{ K}}{0.600 \text{ L}} = 5.7 \times 10^{-4} \text{ atm} = 0.43 \text{ mm Hg}$$

$$10.104 \quad (\text{a}) \quad \text{Change mass } \text{CO}_2 \text{ to mol } \text{CO}_2. \quad P = 1.00 \text{ atm}, T = 27^\circ\text{C} = 300 \text{ K}.$$

$$6 \times 10^6 \text{ tons } \text{CO}_2 \times \frac{2000 \text{ lb}}{\text{ton}} \times \frac{453.6 \text{ g}}{\text{lb}} \times \frac{1 \text{ mol } \text{CO}_2}{44.01 \text{ g } \text{CO}_2} = 1.237 \times 10^{11} = 1 \times 10^{11} \text{ mol}$$

$$V = \frac{nRT}{P} = 1.237 \times 10^{11} \text{ mol} \times \frac{300 \text{ K}}{1.00 \text{ atm}} \times \frac{0.08206 \text{ L-atm}}{\text{mol-K}} = 3.045 \times 10^{12} = 3 \times 10^{12} \text{ L}$$

$$(\text{b}) \quad 1.237 \times 10^{11} \text{ mol } \text{CO}_2 \times \frac{44.01 \text{ g } \text{CO}_2}{\text{mol } \text{CO}_2} \times \frac{1 \text{ cm}^3}{1.2 \text{ g}} \times \frac{1 \text{ L}}{1000 \text{ cm}^3} = 4.536 \times 10^9 = 5 \times 10^9 \text{ L}$$

$$(\text{c}) \quad n = 1.237 \times 10^{11} \text{ mol}, P = 90 \text{ atm}, T = 36^\circ\text{C} = 309 \text{ K}$$

$$V = \frac{nRT}{P} = 1.237 \times 10^{11} \text{ mol} \times \frac{309 \text{ K}}{90 \text{ atm}} \times \frac{0.08206 \text{ L-atm}}{\text{mol-K}} = 3.485 \times 10^{10} = 3 \times 10^{10} \text{ L}$$

$$10.105 \quad (\text{a}) \quad n = \frac{PV}{RT} = 3.00 \text{ atm} \times \frac{\text{mol-K}}{0.08206 \text{ L-atm}} \times \frac{110 \text{ L}}{300 \text{ K}} = 13.4 \text{ mol } \text{C}_3\text{H}_8(\text{g})$$

$$(\text{b}) \quad \frac{0.590 \text{ g } \text{C}_3\text{H}_8(\text{l})}{1 \text{ mL}} \times 110 \times 10^3 \text{ mL} \times \frac{1 \text{ mol } \text{C}_3\text{H}_8}{44.094 \text{ g}} = 1.47 \times 10^3 \text{ mol } \text{C}_3\text{H}_8(\text{l})$$

$$(\text{c}) \quad \text{Using } \text{C}_3\text{H}_8 \text{ in a } 110 \text{ L container as an example, the ratio of moles liquid to moles gas that can be stored in a certain volume is } \frac{1.47 \times 10^3 \text{ mol liquid}}{13.4 \text{ mol gas}} = 110.$$

A container with a fixed volume holds many more moles (molecules) of $\text{C}_3\text{H}_8(\text{l})$ because in the liquid phase the molecules are touching. In the gas phase, the molecules are far apart (statement 2, section 10.7), and many fewer molecules will fit in the container.

$$10.106 \quad \text{Vol of room} = 12 \text{ ft} \times 20 \text{ ft} \times 9 \text{ ft} \times \frac{12^3 \text{ in}^3}{1 \text{ ft}^3} \times \frac{2.54^3 \text{ cm}^3}{1^3 \text{ in}^3} \times \frac{1 \text{ L}}{1000 \text{ cm}^3} = 61,164 = 6 \times 10^4 \text{ L}$$

Calculate the total moles of gas in the laboratory at the conditions given.

$$n_t = \frac{PV}{RT} = 1.00 \text{ atm} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{61,164 \text{ L}}{297 \text{ K}} = 2510 = 3 \times 10^3 \text{ mol gas}$$

A $\text{Ni}(\text{CO})_4$ concentration of 1 part in 10^9 means 1 mol $\text{Ni}(\text{CO})_4$ in 1×10^9 total moles of gas.

$$\frac{x \text{ mol Ni}(\text{CO})_4}{2.510 \times 10^3 \text{ mol gas}} = \frac{1}{1 \times 10^9} = 2.510 \times 10^{-6} = 3 \times 10^{-6} \text{ mol Ni}(\text{CO})_4$$

$$2.510 \times 10^{-6} \text{ mol Ni}(\text{CO})_4 \times \frac{170.74 \text{ g Ni}(\text{CO})_4}{1 \text{ mol Ni}(\text{CO})_4} = 4.286 \times 10^{-3} = 4 \times 10^{-3} \text{ g} = 4 \text{ mg Ni}(\text{CO})_4$$

10.107 (a) mol = g/MM; assume mol Ar = mol X;

$$\frac{\text{g Ar}}{39.948 \text{ g/mol}} = \frac{\text{g X}}{\text{MM X}}; \quad \frac{3.224 \text{ g Ar}}{39.948 \text{ g/mol}} = \frac{8.102 \text{ g X}}{\text{MM X}}$$

$$\text{MM X} = \frac{(8.102 \text{ g X})(39.948 \text{ g/mol})}{3.224 \text{ g Ar}} = 100.39 = 100.4 \text{ g/mol}$$

(b) Assume mol Ar = mol X. For gases, $PV = nRT$ and $n = PV/RT$. For moles of the two gases to be equal, the implied assumption is that P , V , and T are constant. Since we use the same container for both gas samples, constant V is a good assumption. Constant P and T are not explicitly stated.

We also assume that the gases behave ideally. At ambient conditions, this is a reasonable assumption.

10.108 It is simplest to calculate the partial pressure of each gas as it expands into the total volume, then sum the partial pressures.

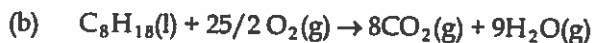
$$P_2 = P_1 V_1 / V_2; \quad P_{\text{N}_2} = 265 \text{ torr} (1.0 \text{ L} / 2.5 \text{ L}) = 106 = 1.1 \times 10^2 \text{ torr}$$

$$P_{\text{Ne}} = 800 \text{ torr} (1.0 \text{ L} / 2.5 \text{ L}) = 320 = 3.2 \times 10^2 \text{ torr}; \quad P_{\text{H}_2} = 532 \text{ torr} (0.5 \text{ L} / 2.5 \text{ L}) \\ = 106 = 1.1 \times 10^2 \text{ torr}$$

$$P_t = P_{\text{N}_2} + P_{\text{Ne}} + P_{\text{H}_2} = (106 + 320 + 106) \text{ torr} = 532 = 5.3 \times 10^2 \text{ torr}$$

$$10.109 \quad (a) \quad n = \frac{PV}{RT} = 0.980 \text{ atm} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{0.524 \text{ L}}{347 \text{ K}} = 0.018034 = 0.0180 \text{ mol air}$$

$$\text{mol O}_2 = 0.018034 \text{ mol air} \times \frac{0.2095 \text{ mol O}_2}{1 \text{ mol air}} = 0.003778 = 0.00378 \text{ mol O}_2$$



(The H_2O produced in an automobile engine is in the gaseous state.)

$$0.003778 \text{ mol O}_2 \times \frac{1 \text{ mol C}_8\text{H}_{18}}{12.5 \text{ mol O}_2} \times \frac{114.2 \text{ g C}_8\text{H}_{18}}{1 \text{ mol C}_8\text{H}_{18}} = 0.0345 \text{ g C}_8\text{H}_{18}$$

- 10.110 (a) Pressure percent = mol percent. Change pressure/mol percents to mol fraction. Partial pressure of each gas is mol fraction (χ) times total pressure. $P_x = \chi_x P_t$

$$P_{N_2} = 0.748(0.985 \text{ atm}) = 0.737 \text{ atm}; \quad P_{O_2} = 0.153(0.985 \text{ atm}) = 0.151 \text{ atm}$$

$$P_{CO_2} = 0.037(0.985 \text{ atm}) = 0.03645 = 0.036 \text{ atm}$$

$$P_{H_2O} = 0.062(0.985 \text{ atm}) = 0.06107 = 0.061 \text{ atm}$$

- (b) $PV = nRT$, $n = PV/RT$; $P = 0.036 \text{ atm}$, $V = 0.455 \text{ L}$, $T = 37^\circ\text{C} = 310 \text{ K}$

$$n = 0.03645 \text{ atm} \times \frac{0.455 \text{ L}}{310 \text{ K}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} = 6.520 \times 10^{-4} = 6.5 \times 10^{-4} \text{ mol}$$

- (c) $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$

$$6.520 \times 10^{-4} \text{ mol } CO_2 \times \frac{1 \text{ mol } C_6H_{12}O_6}{6 \text{ mol } CO_2} \times \frac{180.15 \text{ g } C_6H_{12}O_6}{1 \text{ mol } C_6H_{12}O_6} = 0.01958$$

$$= 0.020 \text{ g } C_6H_{12}O_6$$

- 10.111 V and T are the same for He and O_2 .

$$P_{He} V = n_{He} RT, \quad P_{He} / n_{He} = RT/V; \quad P_{O_2} / n_{O_2} = RT/V$$

$$\frac{P_{He}}{n_{He}} = \frac{P_{O_2}}{n_{O_2}} = n_{O_2} \times \frac{P_{O_2} \times n_{He}}{P_{He}}; \quad n_{He} = 1.42 \text{ g He} \times \frac{1 \text{ mol He}}{4.003 \text{ g He}} = 0.3547 = 0.355 \text{ mol He}$$

$$n_{O_2} = \frac{158 \text{ torr}}{42.5 \text{ torr}} \times 0.355 \text{ mol} = 1.3188 = 1.32 \text{ mol } O_2; \quad 1.3188 \text{ mol } O_2 \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} = 42.2 \text{ g } O_2$$

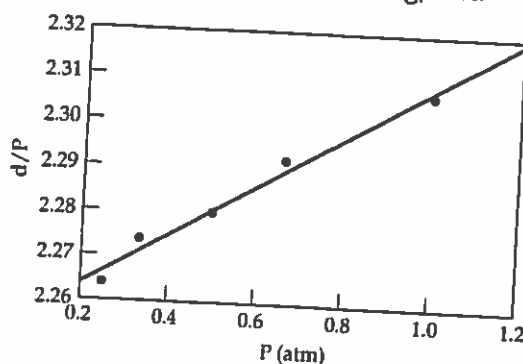
$$10.112 \quad MM_{\text{avg}} = \frac{dRT}{P} = \frac{1.104 \text{ g}}{1 \text{ L}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{400 \text{ K}}{355 \text{ torr}} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 77.579 = 77.6 \text{ g/mol}$$

χ = mole fraction O_2 ; $1 - \chi$ = mole fraction Kr

$$77.579 \text{ g} = \chi(32.00) + (1 - \chi)(83.80)$$

$$6.22 = 51.80 \chi; \quad \chi = 0.120; \quad 12.0\% O_2$$

- 10.113 (a) The quantity $d/P = MM/RT$ should be a constant at all pressures for an ideal gas. It is not, however, because of nonideal behavior. If we graph d/P vs P , the ratio should approach ideal behavior at low P . At $P = 0$, $d/P = 2.2525$. Using this value in the formula $MM = d/P \times RT$, $MM = 50.46 \text{ g/mol}$.



- (b) The ratio d/P varies with pressure because of the finite volumes of gas molecules and attractive intermolecular forces.

- 10.114 Calculate the number of moles of Ar in the vessel:

$$n = (339.854 - 337.428)/39.948 = 0.060729 = 0.06073 \text{ mol}$$

The total number of moles of the mixed gas is the same (Avogadro's Law). Thus, the average atomic weight is $(339.076 - 337.428)/0.060729 = 27.137 = 27.14$. Let the mole fraction of Ne be χ . Then,

$$\chi(20.183) + (1 - \chi)(39.948) = 27.137; 12.811 = 19.765\chi; \chi = 0.6482$$

Neon is thus 64.82 mole percent of the mixture.

- 10.115 $u_{\text{rms}} = (3RT/\text{MM})^{1/2}$; $u_{\text{rms}2} = 2 u_{\text{rms}1}$; $T_1 = -33^\circ = 240 \text{ K}$

$$u_{\text{rms}1} = (3RT_1/\text{MM})^{1/2}; u_{\text{rms}1}^2 = 3(240)R/\text{MM} = 720R/\text{MM}$$

$$u_{\text{rms}1} = (720R/\text{MM})^{1/2}; u_{\text{rms}2} = 2u_{\text{rms}1} = (2)(720R/\text{MM})^{1/2}$$

$$(2)(720R/\text{MM})^{1/2} = (3RT_2/\text{MM})^{1/2}$$

$$(2)^2(720R/\text{MM}) = 3RT_2/\text{MM}; (2)^2(720) = 3T_2$$

$$T_2 = (4)(720)/3 = 960 \text{ K} = 687^\circ\text{C}$$

Increasing the rms speed (u) by a factor of 2 requires heating to 960 K (or 687°C), increasing the temperature by a factor of 4.

- 10.116 (a) Assumption 3 states that attractive and repulsive forces between molecules are negligible. All gases in the list are nonpolar. The largest and most structurally complex molecule, SF_6 , is most likely to depart from this assumption.
- (b) The monatomic gas Ne is smallest and least structurally complex, so it will behave most like an ideal gas.
- (c) Root-mean-square speed is inversely related to molecular mass. The lightest gas, CH_4 , has the highest rms speed.
- (d) The heaviest and most structurally complex is SF_6 . Also, S and F have larger atomic radii than C and H; this means that S-F bonds will be longer than C-H bonds and the volume of SF_6 will be greater than that of CH_4 . It is reasonable to assume that SF_6 will occupy the greatest molecular volume relative to total volume. A quantitative measure is the b value in Table 10.3, with units of L/mol. Unfortunately, SF_6 does not appear in Table 10.3.
- (e) Average kinetic energy is only related to absolute (K) temperature. At the same temperature, they all have the same average kinetic molecular energy.
- (f) Rate of effusion is inversely related to molecular mass. The lighter the molecule, the faster it effuses. Ne and CH_4 have smaller molecular masses and effuse faster than N_2 .
- (g) If SF_6 occupies the greatest molecular volume [see part (d)], we expect it to have the largest van der Waals b parameter.

- 10.117 (a) The effect of intermolecular attraction becomes more significant as a gas is compressed to a smaller volume at constant temperature. This compression causes the pressure, and thus the number of intermolecular collisions, to increase. Intermolecular attraction causes some of these collisions to be inelastic, which amplifies the deviation from ideal behavior.
- (b) The effect of intermolecular attraction becomes less significant as the temperature of a gas is increased at constant volume. When the temperature of a gas is increased at constant volume, the pressure of the gas, the number of intermolecular collisions, and the average kinetic energy of the gas particles increase. This higher average kinetic energy means that a larger fraction of the molecules has sufficient kinetic energy to overcome intermolecular attractions, even though there are more total collisions. This increases the fraction of elastic collisions, and the gas more closely obeys the ideal-gas equation.
- 10.118 The larger and heavier the particle, in this case a single atom, the more likely it is to deviate from ideal behavior. Other than Rn, Xe is the largest (atomic radius = 1.30 Å), heaviest (molar mass = 131.3 g/mol) and most dense (5.90 g/L) noble gas. Its susceptibility to intermolecular interactions is also demonstrated by its ability to form compounds like XeF₄.
- 10.119 (a) At STP, the molar volume = $1 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{273 \text{ K}}{1 \text{ atm}} = 22.4 \text{ L}$
 Dividing the value for *b*, 0.0322 L/mol, by 4, we obtain 0.00805 L. Thus, the volume of the Ar atoms is $(0.00805/22.4)100 = 0.0359\%$ of the total volume.
- (b) At 200 atm pressure (and 0°C, standard temperature) the molar volume is 0.112 L, and the volume of the Ar atoms is 7.19% of the total volume.
- 10.120 (a) $120.00 \text{ kg N}_2(\text{g}) \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol N}_2}{28.0135 \text{ g N}} = 4283.6 \text{ mol N}_2$
 $P = \frac{nRT}{V} = 4283.6 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{553 \text{ K}}{1100.0 \text{ L}} = 176.72 = 177 \text{ atm}$
- (b) According to Equation [10.26], $P = \frac{nRT}{V - nb} - \frac{n^2 a}{V^2}$
 $P = \frac{(4283.6 \text{ mol})(0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(553 \text{ K})}{1100.0 \text{ L} - (4283.6 \text{ mol})(0.0391 \text{ L/mol})} - \frac{(4283.6 \text{ mol})^2 (1.39 \text{ L}^2 \cdot \text{atm/mol}^2)}{(1100.0 \text{ L})^2}$
 $P = \frac{194,388 \text{ L} \cdot \text{atm}}{1100.0 \text{ L} - 167.5 \text{ L}} - 21.1 \text{ atm} = 208.5 \text{ atm} - 21.1 \text{ atm} = 187.4 \text{ atm}$
- (c) The pressure corrected for the real volume of the N₂ molecules is 208.5 atm, 31.8 atm higher than the ideal pressure of 176.7 atm. The 21.1 atm correction for intermolecular forces reduces the calculated pressure somewhat, but the "real" pressure is still higher than the ideal pressure. The correction for the real volume of molecules dominates. Even though the value of *b* is small, the number of moles of N₂ is large enough so that the molecular volume correction is larger than the attractive forces correction.

Integrative Exercises

$$10.121 \quad (a) \quad MM = \frac{gRT}{VP} = \frac{1.56 \text{ g}}{1.00 \text{ L}} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{323 \text{ K}}{0.984 \text{ atm}} = 42.0 \text{ g/mol}$$

Assume 100 g cyclopropane

$$100 \text{ g} \times 0.857 \text{ C} = 85.7 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g}} = \frac{7.136 \text{ mol C}}{7.136} = 1 \text{ mol C}$$

$$100 \text{ g} \times 0.143 \text{ H} = 14.3 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g}} = \frac{14.19 \text{ mol H}}{7.136} = 2 \text{ mol H}$$

The empirical formula of cyclopropane is CH_2 and the empirical formula weight is $12 + 2 = 14 \text{ g}$. The ratio of molar mass to empirical formula weight, $42.0 \text{ g}/14 \text{ g}$, is 3; therefore, there are three empirical formula units in one cyclopropane molecule. The molecular formula is $3 \times (\text{CH}_2) = \text{C}_3\text{H}_6$.

- (b) Ar is a monoatomic gas. Cyclopropane molecules are larger and more structurally complex, even though the molar masses of Ar and C_3H_6 are similar. If both gases are at the same relatively low temperature, they have approximately the same average kinetic energy, and the same ability to overcome intermolecular attractions. We expect intermolecular attractions to be more significant for the more complex C_3H_6 molecules, and that C_3H_6 will deviate more from ideal behavior at the conditions listed. This conclusion is supported by the a values in Table 10.3. The a values for CH_4 and CO_2 , more complex molecules than Ar atoms, are larger than the value for Ar. If the pressure is high enough for the volume correction in the van der Waals equation to dominate behavior, the larger C_3H_6 molecules definitely deviate more than Ar atoms from ideal behavior.
- (c) Cyclopropane, C_3H_6 , $MM = 42.0 \text{ g/mol}$; methane, CH_4 , $MM = 16.0$. Rate of effusion through a pinhole is inversely related to molar mass. Cyclopropane would effuse through a pinhole slower than methane, because it has the greater molar mass.

- 10.122 *Plan.* Write the balanced equation for the combustion of methanol. Since amounts of both reactants are given, determine the limiting reactant. Use mole ratios to calculate mol H_2O produced, based on the amount of limiting reactant. Change moles to g H_2O , then use density to calculate volume of $\text{H}_2\text{O}(\text{l})$ produced. Assume the condensed $\text{H}_2\text{O}(\text{l})$ is at 25°C , where density = 0.99707 g/mol . *Solve.*



$$25.0 \text{ mL CH}_3\text{OH} \times \frac{0.850 \text{ g CH}_3\text{OH}}{\text{mL}} \times \frac{1 \text{ mol CH}_3\text{OH}}{32.04 \text{ g}} = 0.6632 = 0.663 \text{ mol CH}_3\text{OH}$$

$$\text{mol O}_2 = n = \frac{PV}{RT} = 1.00 \text{ atm} \times \frac{12.5 \text{ L}}{273 \text{ K}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} = 0.5580 = 0.558 \text{ mol O}_2$$

$$0.558 \text{ mol O}_2 \times \frac{2 \text{ mol CH}_3\text{OH}}{3 \text{ mol O}_2} = 0.372 \text{ mol CH}_3\text{OH}$$

0.558 mol O_2 can react with only $0.372 \text{ mol CH}_3\text{OH}$, so O_2 is the limiting reactant. Note that a large volume of $\text{O}_2(\text{g})$ is required to completely react with a relatively small volume of $\text{CH}_3\text{OH}(\text{l})$.

$$0.558 \text{ mol O}_2 \times \frac{4 \text{ mol H}_2\text{O}}{3 \text{ mol O}_2} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \times \frac{1 \text{ mL H}_2\text{O}}{0.99707 \text{ g H}_2\text{O}} = 13.446 = 13.4 \text{ mL H}_2\text{O}$$

- 10.123 (a) Get g C from mL CO₂; get g H from mL H₂O. Also calculate mol C and H, to use in part (b). Get g N by subtraction. Calculate % composition. $n = PV/RT$. At STP, $P = 1 \text{ atm}$, $T = 273 \text{ K}$. (STP implies an infinite number of sig figs)

$$n_{\text{CO}_2} = 0.08316 \text{ L} \times \frac{1 \text{ atm}}{273 \text{ K}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} = 0.003712 \text{ mol CO}_2$$

$$0.003712 \text{ mol CO}_2 \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \times \frac{12.0107 \text{ g C}}{\text{mol C}} = 0.044585 = 0.04458 \text{ g CO}_2$$

$$n_{\text{H}_2\text{O}} = 0.07330 \text{ L} \times \frac{1 \text{ atm}}{273 \text{ K}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} = 3.2720 \times 10^{-3}$$

$$= 3.272 \times 10^{-3} \text{ mol H}_2\text{O}$$

$$3.2720 \times 10^{-3} \text{ mol H}_2\text{O} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \times \frac{1.00794 \text{ g H}}{\text{mol H}} = 6.5959 \times 10^{-3}$$

$$= 6.596 \times 10^{-3} \text{ g H}$$

$$\text{mass \% X} = \frac{\text{mass X}}{\text{sample mass}} \times 100; \text{ sample mass} = 100.0 \text{ mg} = 0.1000 \text{ g}$$

$$\% \text{ C} = \frac{0.044585 \text{ g}}{0.1000 \text{ g}} \times 100 = 44.585 = 44.58\% \text{ C}$$

$$\% \text{ H} = \frac{6.5959 \times 10^{-3} \text{ g H}}{0.1000 \text{ g}} \times 100 = 6.5959 = 6.596\% \text{ H}$$

$$\% \text{ Cl} = \frac{0.01644 \text{ g Cl}}{0.1000 \text{ g}} \times 100 = 16.44\% \text{ Cl}$$

$$\% \text{ N} = 100 - 44.58 - 6.596 - 16.44 = 32.38\% \text{ N}$$

- (b) 0.003712 mol C ; $2(3.272 \times 10^{-3}) = 6.544 \times 10^{-3} \text{ mol H}$

$$0.01644 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.453 \text{ g Cl}} = 4.637 \times 10^{-4} \text{ mol Cl}$$

$$0.1000 \text{ g sample} \times 0.3238 \text{ mass fraction N} = 0.03238 \text{ g N}$$

$$0.03238 \text{ g N} \times \frac{1 \text{ mol N}}{14.0067 \text{ g N}} = 0.0023118 = 0.002312 \text{ mol N}$$

Divide by the smallest number of mol to find the simplest ratio of moles.

$$\frac{0.003712 \text{ mol C}}{4.637 \times 10^{-4}} = 8.005 \text{ C}$$

$$\frac{6.544 \times 10^{-3} \text{ mol H}}{4.637 \times 10^{-4}} = 14.11 \text{ H}$$

$$\frac{4.637 \times 10^{-4} \text{ mol Cl}}{4.637 \times 10^{-4}} = 1.000 \text{ Cl}$$

$$\frac{0.002312 \text{ mol N}}{4.637 \times 10^{-4}} = 4.985 \text{ N}$$

If we assume 14.11 is "close" to 14 (a reasonable assumption), the empirical formula is $C_3H_{14}N_5Cl$.

- (c) Molar mass of the compound is required in order to determine molecular formula when the empirical formula is known.

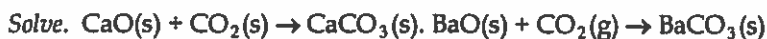
- 10.124 (a) *Plan.* Use the ideal-gas law to calculate the moles CO_2 that react.

Solve. $P(\text{reacted}) = P(\text{initial}) - P(\text{final})$, at constant V, T . Since both CaO and BaO react with CO_2 in a 1:1 mole ratio, $\text{mol } CaO + \text{mol } BaO = \text{mol } CO_2$. Use molar masses to calculate % CaO in sample.

$$P(\text{reacted}) = 730 \text{ torr} - 150 \text{ torr} = 580 \text{ torr}; 580 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.76316 = 0.763 \text{ atm}$$

$$n = \frac{PV}{RT} = 0.76316 \text{ atm} \times \frac{1.0 \text{ L}}{298 \text{ K}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} = 0.03121 = 0.0312 \text{ mol } CO_2$$

- (b) *Plan.* Use the stoichiometry of the reaction and definition of moles to calculate the mass and Mass % of CaO .



$$\text{mol } CO_2 \text{ reacted} = \text{mol } CaO + \text{mol } BaO$$

$$\text{Let } x = \text{g } CaO, 4.00 - x = \text{g } BaO$$

$$0.03121 = \frac{x}{56.08} + \frac{4.00 - x}{153.3}$$

$$0.03121(56.08)(153.3) = 153.3x + 56.08(4.00 - x)$$

$$268.3 = (153.3x - 56.08x) + 224.3$$

$$43.98 = 97.22x, x = 0.452 = 0.45 \text{ g } CaO$$

$$\frac{0.452 \text{ g } CaO}{4.00 \text{ g sample}} \times 100 = 11.3 = 11\% \text{ } CaO$$

(By strict sig fig rules, the result has 2 sig figs, because $268 - 224 = 44$ has 0 decimal places and 2 sig figs.)

- 10.125 (a) $5.00 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} = 0.1371 = 0.137 \text{ mol HCl}$

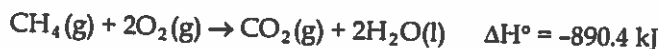
$$5.00 \text{ g } NH_3 \times \frac{1 \text{ mol } NH_3}{17.03 \text{ g } NH_3} = 0.2936 = 0.294 \text{ mol } NH_3$$

The gases react in a 1:1 mole ratio, HCl is the limiting reactant and is completely consumed. $(0.2936 \text{ mol} - 0.1371 \text{ mol}) = 0.1565 = 0.157 \text{ mol } NH_3$ remain in the system. $NH_3(g)$ is the only gas remaining after reaction.

- (b) $V_1 = 4.00 \text{ L}$. $P = \frac{nRT}{V} = 0.1565 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{298 \text{ K}}{4.00 \text{ L}} = 0.957 \text{ atm}$

- (c) $0.137 \text{ mol HCl} \times \frac{1 \text{ mol } NH_4Cl}{1 \text{ mol HCl}} \times \frac{53.49 \text{ g } NH_4Cl}{1 \text{ mol } NH_4Cl} = 7.3284 = 7.33 \text{ g } NH_4Cl$

$$10.126 \quad n = \frac{PV}{RT} = 1.00 \text{ atm} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{2.7 \times 10^{12} \text{ L}}{273 \text{ K}} = 1.205 \times 10^{11} = 1.2 \times 10^{11} \text{ mol CH}_4$$



(At STP, H₂O is in the liquid state.)

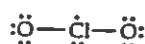
$$\Delta H_{\text{rxn}}^\circ = \Delta H_f^\circ \text{CO}_2(\text{g}) + 2\Delta H_f^\circ \text{H}_2\text{O}(\text{l}) - \Delta H_f^\circ \text{CH}_4(\text{g}) - 2\Delta H_f^\circ \text{O}_2(\text{g})$$

$$\Delta H_{\text{rxn}}^\circ = -393.5 \text{ kJ} + 2(-285.83 \text{ kJ}) - (-74.8 \text{ kJ}) - 0 = -890.4 \text{ kJ}$$

$$\frac{-890.4 \text{ kJ}}{1 \text{ mol CH}_4} \times 1.205 \times 10^{11} \text{ mol CH}_4 = -1.073 \times 10^{14} = -1.1 \times 10^{14} \text{ kJ}$$

The negative sign indicates heat evolved by the combustion reaction.

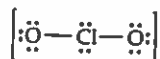
- 10.127 (a) 19 e⁻, 9.5 e⁻ pairs



Resonance structures can be drawn with the odd electron on O, but electronegativity considerations predict that it will be on Cl for most of the time.

- (b) ClO₂ is very reactive because it is an odd-electron molecule. Adding an electron (reduction) both pairs the odd electron and completes the octet of Cl. Thus, ClO₂ has a strong tendency to gain an electron and be reduced.

- (c) ClO₂⁻, 20 e⁻, 10 e⁻ pairs



- (d) 4 e⁻ domains around Cl, O-Cl-O bond angle ~107° (<109° owing to repulsion by nonbonding domains)
- (e) Calculate mol Cl₂ from ideal-gas equation; determine limiting reactant; mass ClO₂ via mol ratios.

$$\text{mol Cl}_2 = \frac{PV}{RT} = 1.50 \text{ atm} \times \frac{2.00 \text{ L}}{294 \text{ K}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} = 0.1243 = 0.124 \text{ mol Cl}_2$$

$$15.0 \text{ g NaClO}_2 \times \frac{1 \text{ mol NaClO}_2}{90.44 \text{ g}} = 0.1659 = 0.166 \text{ mol NaClO}_2$$

2 mol NaClO₂ are required for 1 mol Cl₂, so NaClO₂ is the limiting reactant. For every 2 mol NaClO₂ reacted, 2 mol ClO₂ are produced, so mol ClO₂ = mol NaClO₂.

$$0.1659 \text{ mol ClO}_2 \times \frac{67.45 \text{ g ClO}_2}{\text{mol}} = 11.2 \text{ g ClO}_2$$

- 10.128 (a) ft³ CH₄ → L CH₄ → mol CH₄ → mol CH₃OH → g CH₃OH → L CH₃OH

$$10.7 \times 10^9 \text{ ft}^3 \text{ CH}_4 \times \frac{1 \text{ yd}^3}{3^3 \text{ ft}^3} \times \frac{1 \text{ m}^3}{(1.0936)^3 \text{ yd}^3} \times \frac{1 \text{ L}}{1 \times 10^{-3} \text{ m}^3} = 3.03001 \times 10^{11}$$

$$= 3.03 \times 10^{11} \text{ L CH}_4$$

$$n = \frac{PV}{RT} = \frac{3.03 \times 10^{11} \text{ L} \times 1.00 \text{ atm}}{298 \text{ K}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} = 1.2391 \times 10^{10}$$

$$= 1.24 \times 10^{10} \text{ mol CH}_4$$

$$1 \text{ mol CH}_4 = 1 \text{ mol CH}_3\text{OH}$$

$$1.2391 \times 10^{10} \text{ mol CH}_3\text{OH} \times \frac{32.04 \text{ g CH}_3\text{OH}}{\text{mol CH}_3\text{OH}} \times \frac{1 \text{ mL CH}_3\text{OH}}{0.791 \text{ g}} \times \frac{1 \text{ L}}{1000 \text{ mL}}$$

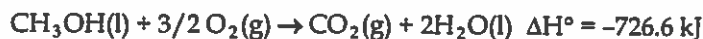
$$= 5.0189 \times 10^8 = 5.02 \times 10^8 \text{ L CH}_3\text{OH}$$



$$\Delta H^\circ = \Delta H_f^\circ \text{CO}_2(\text{g}) + 2\Delta H_f^\circ \text{H}_2\text{O}(\text{l}) - \Delta H_f^\circ \text{CH}_4(\text{g}) - 2\Delta H_f^\circ \text{O}_2(\text{g})$$

$$\Delta H^\circ = -393.5 \text{ kJ} + 2(-285.83 \text{ kJ}) - (-74.8 \text{ kJ}) - 0 = -890.4 \text{ kJ}$$

$$1.2391 \times 10^{10} \text{ mol CH}_4 \times \frac{-890.4 \text{ kJ}}{1 \text{ mol CH}_4} = -1.10 \times 10^{13} \text{ kJ}$$



$$\Delta H^\circ = \Delta H_f^\circ \text{CO}_2(\text{g}) + 2\Delta H_f^\circ \text{H}_2\text{O}(\text{l}) - \Delta H_f^\circ \text{CH}_3\text{OH}(\text{l}) - 3/2 \Delta H_f^\circ \text{O}_2(\text{g})$$

$$= -393.5 \text{ kJ} + 2(-285.83 \text{ kJ}) - (-238.6 \text{ kJ}) - 0 = -726.6 \text{ kJ}$$

$$1.2391 \times 10^{10} \text{ mol CH}_3\text{OH} \times \frac{-726.6 \text{ kJ}}{1 \text{ mol CH}_3\text{OH}} = -9.00 \times 10^{12} \text{ kJ}$$

(c) Assume a volume of 1.00 L of each liquid.

$$1.00 \text{ L CH}_4(\text{l}) \times \frac{466 \text{ g}}{1 \text{ L}} \times \frac{1 \text{ mol}}{16.04 \text{ g}} \times \frac{-890.4 \text{ kJ}}{\text{mol CH}_4} = -2.59 \times 10^4 \text{ kJ/L CH}_4$$

$$1.00 \text{ L CH}_3\text{OH} \times \frac{791 \text{ g}}{1 \text{ L}} \times \frac{1 \text{ mol}}{32.04 \text{ g}} \times \frac{-726.6 \text{ kJ}}{\text{mol CH}_3\text{OH}} = -1.79 \times 10^4 \text{ kJ/L CH}_3\text{OH}$$

Clearly $\text{CH}_4(\text{l})$ has the higher enthalpy of combustion per unit volume.

10.129 After reaction, the flask contains $\text{IF}_5(\text{g})$ and whichever reactant is in excess. Determine the limiting reactant, which regulates the moles of IF_5 produced and moles of excess reactant.



$$10.0 \text{ g I}_2 \times \frac{1 \text{ mol I}_2}{253.8 \text{ g I}_2} \times \frac{5 \text{ mol F}_2}{1 \text{ mol I}_2} = 0.1970 = 0.197 \text{ mol F}_2$$

$$10.0 \text{ g F}_2 \times \frac{1 \text{ mol F}_2}{38.00 \text{ g F}_2} = 0.2632 = 0.263 \text{ mol F}_2 \text{ available}$$

I_2 is the limiting reactant; F_2 is in excess.

$$0.263 \text{ mol F}_2 \text{ available} - 0.197 \text{ mol F}_2 \text{ reacted} = 0.066 \text{ mol F}_2 \text{ remain.}$$

$$10.0 \text{ g I}_2 \times \frac{1 \text{ mol I}_2}{253.8 \text{ g I}_2} \times \frac{2 \text{ mol IF}_5}{1 \text{ mol I}_2} = 0.0788 \text{ mol IF}_5 \text{ produced}$$

$$(a) \quad P_{\text{IF}_5} = \frac{nRT}{V} = 0.0788 \text{ mol} \times \frac{0.08206 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times \frac{398 \text{ K}}{5.00 \text{ L}} = 0.515 \text{ atm}$$

$$(b) \quad \chi_{\text{IF}_5} = \frac{\text{mol IF}_5}{\text{mol IF}_5 + \text{mol F}_2} = \frac{0.0788}{0.0788 + 0.066} = 0.544$$

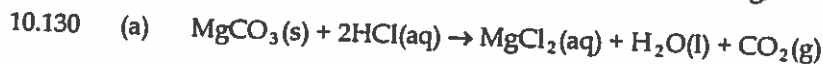
(c) 42 valence e⁻, 21 e⁻ pairs



$$(d) \quad 0.0788 \text{ mol IF}_5 \times \frac{221.90 \text{ g IF}_5}{\text{mol IF}_5} = 17.4857 = 17.5 \text{ g IF}_5 \text{ produced}$$

$$0.066 \text{ mol F}_2 \times \frac{38.00 \text{ g F}_2}{\text{mol F}_2} = 2.508 = 2.5 \text{ g F}_2 \text{ remain}$$

Total mass in flask = 17.5 g IF₅ + 2.5 g F₂ = 20.00 g; mass is conserved.



$$(b) \quad n = \frac{PV}{RT} = 743 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} \times \frac{\text{mol} \cdot \text{K}}{0.08206 \text{ L} \cdot \text{atm}} \times \frac{1.72 \text{ L}}{301 \text{ K}}$$

$$= 0.06808 = 0.0681 \text{ mol CO}_2$$

(c) $x = \text{g MgCO}_3$, $y = \text{g CaCO}_3$, $x + y = 6.53 \text{ g}$
 $\text{mol MgCO}_3 + \text{mol CaCO}_3 = \text{mol CO}_2 \text{ total}$

$$\frac{x}{84.32} + \frac{y}{100.09} = 0.06808; y = 6.53 - x$$

$$\frac{x}{84.32} + \frac{6.53 - x}{100.09} = 0.06808$$

$$100.09x - 84.32x + 84.32(6.53) = 0.06808(84.32)(100.09)$$

$$15.77x + 550.610 = 574.549; x = 1.52 \text{ g MgCO}_3$$

$$\text{mass \% MgCO}_3 = \frac{1.52 \text{ g MgCO}_3}{6.53 \text{ g sample}} \times 100 = 23.3\%$$

[By strict sig fig rules, the answer has 2 sig figs: $15.77x + 551$ (3 digits from 6.53) = 575; $575 - 551 = 24$ (no decimal places, 2 sig figs) leads to 1.5 g MgCO₃ and 23% MgCO₃]

11

Liquids and Intermolecular Forces

Visualizing Concepts

In this chapter we will use the temperature units °C and K interchangeably when designating specific heats and *changes* in temperature.

- 11.1 The diagram best describes a liquid. In the diagram, the particles are close together, mostly touching but there is no regular arrangement or order. This rules out a gaseous sample, where the particles are far apart, and a crystalline solid, which has a regular repeating structure in all three directions.
- 11.2 (a) (i) Hydrogen bonding; H-F interactions qualify for this narrowly defined interaction.
- (ii) London dispersion forces, the only intermolecular forces between nonpolar F₂ molecules.
- (iii) Ion-dipole forces between Na⁺ cation and the negative end of a polar covalent water molecule.
- (iv) Dipole-dipole forces between oppositely charged portions of two polar covalent SO₂ molecules.
- (b) London dispersion forces in (ii) are probably the weakest.
- 11.3 The viscosity of glycerol will be greater than that of 1-propanol. Viscosity is the resistance of a substance to flow. The stronger the intermolecular forces in a liquid, the greater its viscosity. Hydrogen bonding is the predominant force for both molecules. Glycerol has three times as many O-H groups and many more H-bonding interactions than 1-propanol, so it experiences stronger intermolecular forces and greater viscosity. (Both molecules have the same carbon-chain length, so dispersion forces are similar.)
- 11.4 When heat is added to a liquid, the temperature of the liquid rises. If enough heat is added to reach the boiling point, any excess heat is used to vaporize the liquid. If heat is still available when all the liquid is converted to gas, the temperature of the gas rises.

Use the specific heat of CH₄(l) to calculate the amount of heat required to raise the temperature of 32.0 g of CH₄(l) from -170 °C to -161.5 °C. If this is less than 42 kJ, use ΔH_{vap} to calculate the energy required to vaporize the liquid, and so on, until exactly 42.0 kJ has been used to increase the temperature and/or change the state of CH₄.

Heat the liquid to its boiling point: ΔT = [-161.5 °C - (-170 °C)] = 8.5 °C = 8.5 K

$$\frac{3.48 \text{ J}}{\text{g} \cdot \text{K}} \times 32.0 \text{ g CH}_4 \times 8.5 \text{ °C} = 946.56 = 9.5 \times 10^2 \text{ J} = 0.95 \text{ kJ}$$