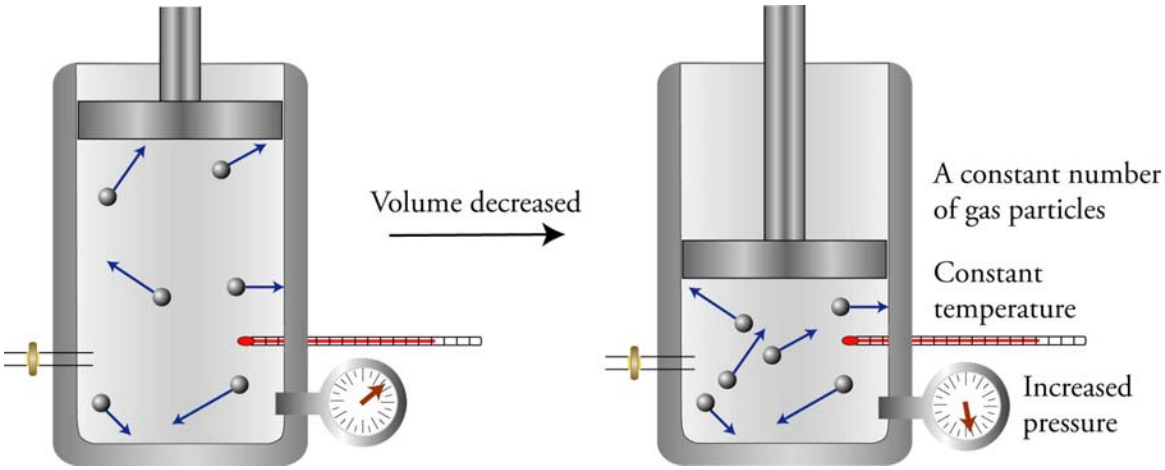
Chapter 11 - Gases



Section Goals and Introductions

Section 11.1 Gases and Their Properties

Goals

- *To describe the particle nature of both real and ideal gases.*
- *To describe the properties of gases that can be used to explain their characteristics: volume, number of particles, temperature, and pressure.*
- *To describe and explain the relationships between the properties of gases.*
- *To use the understanding of the relationships between gas properties to explain real-world things, such as the mechanics of a gasoline engine and the process of breathing.*

This section will increase your understanding of gases by adding more detail to the description of gases first presented in Chapter 3. As you know, it is often useful for scientists and science students to use simplified versions of reality (models) to explain scientific phenomena. This section introduces the ideal gas model that helps us to explain the characteristics of most gases.

The next portion of this section describes the properties of gases (number of gas particles, volume, temperature, and gas pressure) with an emphasis on gas pressure. You will discover what gas pressure is and what causes it. The rest of this section shows how the ideal gas model can be used to explain the relationships between the properties. For example, you will learn why increased temperature leads to increased pressure for a constant amount of gas in a constant volume. Spend the time it takes to develop a mental image of the particle nature of gases, and be sure that you can use that image to *see* how changing one property of a gas leads to changes in others.

Internet: Boyle's Law

Internet: Gay-Lussac's Law

Internet: Charles' Law

Internet: Number of Particles and Pressure

Internet: Avogadro's Law

Section 11.2 Ideal Gas Calculations

Goal: To show how the properties of gases can be calculated.

This section derives two equations that relate to ideal gases (called the ideal gas equation and the combined gas law equation) and shows how you can use these equations to calculate values for gas properties. Pay special attention to the two sample study sheets that will help you to develop logical procedures for these calculations. See the related section on our Web site:

Internet: Real Gases

Section 11.3 Equation Stoichiometry and Ideal Gases

Goal: To show how gas-related calculations can be applied to equation stoichiometry problems.

This section shows how we can combine calculations such as those found in Chapter 10 with the gas calculations described in Section 11.2 to do equation stoichiometry problems that include gaseous reactants and products. It is a good idea to review Study Sheet 10.3: Equation Stoichiometry Problems and perhaps other parts of Chapter 10 before reading this section. Sample Study Sheet 11.3: Equation Stoichiometry Problems summarizes the procedures for equation stoichiometry problems that allow you to convert between amount of one reactant or product and amount of another reactant or product whether these substances are pure solids, pure liquids, pure gases, or in water solutions (aqueous). See the related section on our Web site:

Internet: Gas Stoichiometry Shortcut

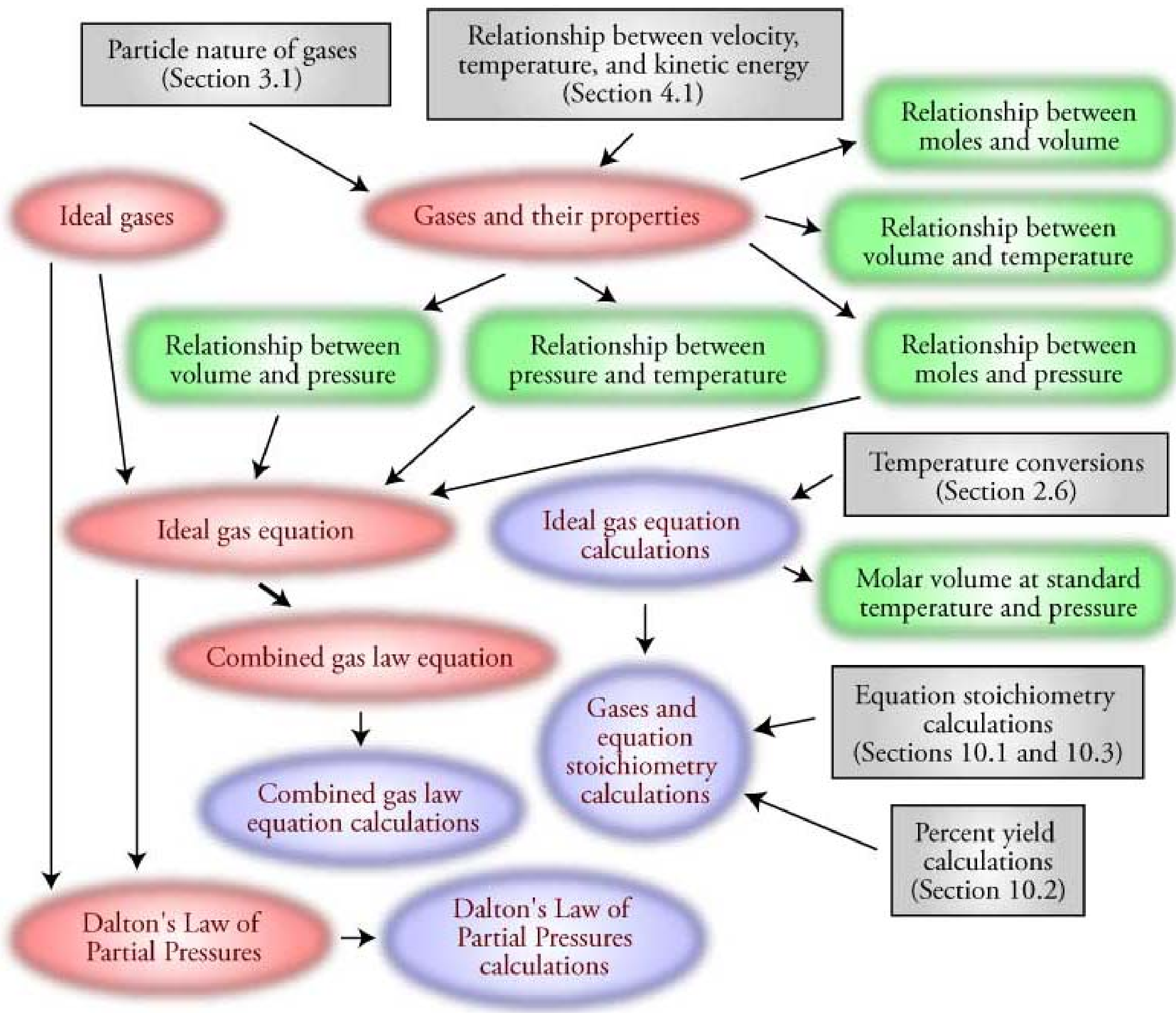
Section 11.4 Dalton's Law of Partial Pressures

Goals

- *To describe the properties of mixtures of gases.*
- *To describe calculations that deal with mixtures of gases.*

In the real world, gases are usually mixtures. This section describes how mixing gases affects the properties of the resulting mixture. Be sure that you can visualize mixtures of gases and that you can use this image to help you understand the effect that mixing gases has on the overall pressure created by the mixture. The section also derives two equations that allow you to calculate the pressures of gaseous mixtures.

Chapter 11 Map



Chapter Checklist

- ☐ Read the Review Skills section. If there is any skill mentioned that you have not yet mastered, review the material on that topic before reading this chapter.
- ☐ Read the chapter quickly before the lecture that describes it.
- ☐ Attend class meetings, take notes, and participate in class discussions.
- ☐ Work the Chapter Exercises, perhaps using the Chapter Examples as guides.
- ☐ Study the Chapter Glossary and test yourself on our Web site:

  Internet: Glossary Quiz
- ☐ Study all of the Chapter Objectives. You might want to write a description of how you will meet each objective.

- ☐ Reread the Study Sheets in this chapter and decide whether you will use them or some variation on them to complete the tasks they describe.
  Sample Study Sheet 11.1: Using the Ideal Gas Equation
  Sample Study Sheet 11.2: Using the Combined Gas Law Equation
  Sample Study Sheet 11.3: Equation Stoichiometry Problems
  Sample Study Sheet 11.4: Using Dalton's Law of Partial Pressures
- ☐ This chapter has logic sequences in Figures 11.3, 11.4, 11.5, 11.6, and 11.7. Convince yourself that each of the statements in these sequences logically lead to the next statement.
- ☐ Memorize the following equations.

$$PV = nRT$$

$$PV = \frac{g}{M}RT$$

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

$$P_{\text{total}} = \sum P_{\text{partial}}$$

$$P_{\text{total}} = \left(\sum n_{\text{each gas}}\right)\frac{RT}{V}$$

- ☐ To get a review of the most important topics in the chapter, fill in the blanks in the Key Ideas section.
- ☐ Work all of the selected problems at the end of the chapter, and check your answers with the solutions provided in this chapter of the study guide.
- ☐ Ask for help if you need it.

Web Resources

Internet: Boyle's Law
Internet: Gay-Lussac's Law
Internet: Charles' Law
Internet: Number of Particles and Pressure
Internet: Avogadro's Law
Internet: Real Gases
Internet: Gas Stoichiometry Shortcut
Internet: Glossary Quiz

Exercise Key

Exercise 11.1 – Using the Ideal Gas Equation: Krypton gas does a better job than argon of slowing the evaporation of the tungsten filament in an incandescent light bulb. Because of its higher cost, however, krypton is used only when longer life is considered to be worth the extra expense. *(Objs 15, 16, & 17)*

a. How many moles of krypton gas must be added to a 175 mL incandescent light bulb to yield a gas pressure of 117 kPa at 21.6 °C?

$$PV = nRT \qquad n = \frac{PV}{RT} = \frac{117\text{ kPa}\,(175\text{ mL})}{\frac{8.3145\text{ L}\bullet\text{kPa}}{\text{K}\bullet\text{mol}}\;294.8\text{ K}}\left(\frac{1\text{ L}}{10^3\text{ mL}}\right) = \mathbf{8.35 \times 10^{-3}\ mol\ Kr}$$

b. What is the volume of an incandescent light bulb that contains 1.196 g of Kr at a pressure of 1.70 atm and a temperature of 97 °C?

$V = ?$ $g = 1.196\text{ g}$ $T = 97\text{ °C} + 273.15 = 370\text{ K}$ $P = 1.70\text{ atm}$

$$PV = \frac{g}{M}RT \qquad V = \frac{gRT}{PM} = \frac{1.196\text{ g Kr}\left(\frac{0.082058\text{ L}\bullet\text{atm}}{\text{K}\bullet\text{mole}}\right)370\text{ K}}{1.70\text{ atm}\left(83.80\,\frac{\text{g}}{\text{mole}}\right)} = \mathbf{0.255\ L\ Kr}$$

c. What is the density of krypton gas at 18.2 °C and 762 mmHg?

$\frac{g}{V} = ?$ $P = 762\text{ mmHg}$ $T = 18.2\text{ °C} + 273.15 = 291.4\text{ K}$

$$PV = \frac{g}{M}RT \qquad \frac{g}{V} = \frac{PM}{RT} = \frac{762\text{ mmHg}\left(83.80\,\frac{\text{g}}{\text{mole}}\right)}{0.082058\,\frac{\text{L}\bullet\text{atm}}{\text{K}\bullet\text{mole}}\,(291.4\text{ K})}\left(\frac{1\text{ atm}}{760\text{ mmHg}}\right) = \mathbf{3.51\ g/L}$$

Exercise 11.2 – Using the Combined Gas Law Equation: A helium weather balloon is filled in Monterey, California, on a day when the atmospheric pressure is 102 kPa and the temperature is 18 °C. Its volume under these conditions is 1.6×10^4 L. Upon being released, it rises to an altitude where the temperature is –8.6 °C, and its volume increases to 4.7×10^4 L. Assuming that the internal pressure of the balloon equals the atmospheric pressure, what is the pressure at this altitude? *(Obj 18)*

$P_1 = 102\text{ kPa}$ $T_1 = 18\text{ °C} + 273.15 = 291\text{ K}$ $V_1 = 1.6 \times 10^4\text{ L}$

$P_2 = ?$ $T_2 = -8.6\text{ °C} + 273.15 = 264.6\text{ K}$ $V_2 = 4.7 \times 10^4\text{ L}$

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} \quad \text{to} \quad \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$P_2 = P_1\left(\frac{T_2}{T_1}\right)\left(\frac{V_1}{V_2}\right) = 102\text{ kPa}\left(\frac{264.6\text{ K}}{291\text{ K}}\right)\left(\frac{1.6\times10^4\text{ L}}{4.7\times10^4\text{ L}}\right) = \mathbf{32\ kPa}$$

Exercise 11.3 – Equation Stoichiometry: Iron is combined with carbon in a series of reactions to form pig iron, which is about 4.3% carbon.

$$2C + O_2 \rightarrow 2CO$$

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

$$2CO \rightarrow C\text{ (in iron)} + CO_2$$

Pig iron is easier to shape than pure iron, and the presence of carbon lowers its melting point from the 1539 °C required to melt pure iron to 1130 °C. *(Obj 19)*

a. In the first reaction, what minimum volume of oxygen at STP is necessary to convert 125 Mg of carbon to carbon monoxide?

$$? \text{ L } O_2 = 125 \text{ Mg C}\left(\frac{10^6 \text{ g}}{1 \text{ Mg}}\right)\left(\frac{1 \text{ mol C}}{12.011 \text{ g C}}\right)\left(\frac{1 \text{ mol } O_2}{2 \text{ mol C}}\right)\left(\frac{22.414 \text{ L } O_2}{1 \text{ mol } O_2}\right)_{\text{STP}}$$

$$\mathbf{= 1.17 \times 10^8 \text{ L } O_2 \quad or \quad 1.17 \times 10^5 \text{ m}^3 \text{ } O_2}$$

b. In the first reaction, what is the maximum volume of carbon monoxide at 1.05 atm and 35 °C that could form from the conversion of 8.74×10^5 L of oxygen at 0.994 atm and 27 °C?

$$? \text{ L CO} = 8.74 \times 10^5 \text{ L } O_2\left(\frac{\text{K} \bullet \text{mol}}{0.082058 \text{ L} \bullet \text{atm}}\right)\left(\frac{0.994 \text{ atm}}{300 \text{ K}}\right)\left(\frac{2 \text{ mol CO}}{1 \text{ mol } O_2}\right)$$

$$\left(\frac{0.082058 \text{ L} \bullet \text{atm}}{\text{K} \bullet \text{mol}}\right)\left(\frac{308 \text{ K}}{1.05 \text{ atm}}\right) = \mathbf{1.70 \times 10^6 \text{ L CO}}$$

Exercise 11.4 – Equation Stoichiometry: Sodium hypochlorite, NaOCl, which is found in household bleaches, can be made from a reaction using chlorine gas and aqueous sodium hydroxide:

$$Cl_2(g) + 2NaOH(aq) \rightarrow NaOCl(aq) + NaCl(aq) + H_2O(l)$$

What minimum volume of chlorine gas at 101.4 kPa and 18.0 °C must be used to react with all the sodium hydroxide in 3525 L of 12.5 M NaOH? *(Obj 19)*

$$? \text{ L } Cl_2 = 3525 \text{ L NaOH soln}\left(\frac{12.5 \text{ mol NaOH}}{1 \text{ L NaOH soln}}\right)\left(\frac{1 \text{ mol } Cl_2}{2 \text{ mol NaOH}}\right)\left(\frac{8.3145 \text{ L} \bullet \text{kPa}}{\text{K} \bullet \text{mol}}\right)\left(\frac{291.0 \text{ K}}{101.4 \text{ kPa}}\right)$$

$$\mathbf{= 5.26 \times 10^5 \text{ L } Cl_2}$$

Exercise 11.5 – Dalton's Law of Partial Pressures: A typical "neon light" contains neon gas mixed with argon gas. (*Objs 21 & 22*)

a. If the total pressure of the mixture of gases is 1.30 kPa and the partial pressure of neon gas is 0.27 kPa, what is the partial pressure of the argon gas?

$$P_{total} = P_{Ne} + P_{Ar} \qquad P_{Ar} = P_{total} - P_{Ne} = 1.30\text{ kPa} - 0.27\text{ kPa} = \mathbf{1.03\ kPa}$$

b. If 6.3 mg of Ar and 1.2 mg Ne are added to the 375-mL tube at 291 K, what is the total pressure of the gases in millimeters of mercury?

$$?\text{ mol Ar} = 6.3\text{ mg Ar}\left(\frac{1\text{ g}}{10^3\text{ mg}}\right)\left(\frac{1\text{ mol Ar}}{39.948\text{ g Ar}}\right) = 1.6 \times 10^{-4}\text{ mol Ar}$$

$$?\text{ mol Ne} = 1.2\text{ mg Ne}\left(\frac{1\text{ g}}{10^3\text{ mg}}\right)\left(\frac{1\text{ mol Ne}}{20.1797\text{ g Ne}}\right) = 5.9 \times 10^{-5}\text{ mol Ne}$$

$$P_{total} = \left(\sum n\right)\frac{RT}{V}$$

$$= \left(1.6\times10^{-4}\text{ mol} + 5.9\times10^{-5}\text{ mol}\right)\frac{\dfrac{0.082058\text{ L}\bullet\text{atm}}{\text{K}\bullet\text{mol}}(291\text{ K})}{375\text{ mL}}\left(\frac{10^3\text{ mL}}{1\text{ L}}\right)\left(\frac{760\text{ mmHg}}{1\text{ atm}}\right)$$

= 11 mmHg

Review Questions Key

1. Describe the particle nature of gases. (See Section 3.1.)

   Our model of the nature of matter has the following components:

   - *All matter is composed of tiny particles.*
   - *These particles are in constant motion. The amount of motion is proportional to temperature. Increased temperature means increased motion.*
   - *Solids, gases, and liquids differ in the degree of motion of their particles and the extent to which the particles interact.*

   Because the particles of a gas are much farther apart than those of the solid or liquid, the particles do not have significant attractions between them. The particles in a gas move freely in straight-line paths until they collide with another particle or the walls of the container. If you were riding on a particle in the gas state, your ride would be generally boring with regular interruptions caused by violent collisions. Between these collisions, you would not even know there were other particles in the container. Because the particles are moving with different velocities and in different directions, the collisions lead to constant changes in the direction and velocity of the motion of each particle. The rapid, random movement of the gas particles allows gases to adjust to the shape and volume of their container.

2. What is 265.2 °C on the Kelvin scale? Convert 565.7 K into °C.

   ? K = 265.2 °C + 273.15 K = **538.4 K** ? °C = 565.7 K – 273.15 = **292.6 °C**

3. About 55% of industrially produced sodium sulfate is used to make detergents. It is made from the reaction

$$4NaCl + 2SO_2 + 2H_2O + O_2 \rightarrow 2Na_2SO_4 + 4HCl$$

a. What is the maximum mass of sodium sulfate that can be produced in the reaction of 745 Mg of sodium chloride with excess SO_2, H_2O, and O_2?

$$? \text{ Mg } Na_2SO_4 = 745 \text{ Mg NaCl}\left(\frac{10^6 \text{ g}}{1 \text{ Mg}}\right)\left(\frac{1 \text{ mol NaCl}}{58.4425 \text{ g NaCl}}\right)\left(\frac{2 \text{ mol } Na_2SO_4}{4 \text{ mol NaCl}}\right)$$
$$\left(\frac{142.043 \text{ g } Na_2SO_4}{1 \text{ mol } Na_2SO_4}\right)\left(\frac{1 \text{ Mg}}{10^6 \text{ g}}\right)$$

or using a shortcut $? \text{ Mg } Na_2SO_4 = 745 \text{ Mg NaCl}\left(\frac{2 \times 142.043 \text{ Mg } Na_2SO_4}{4 \times 58.4425 \text{ Mg NaCl}}\right)$ **= 905 Mg**

b. What is the maximum mass of sodium sulfate that can be produced in the reaction of 745 Mg of sodium chloride with 150 Mg H_2O and excess SO_2 and O_2?

$$? \text{ Mg } Na_2SO_4 = 745 \text{ Mg NaCl}\left(\frac{2 \times 142.043 \text{ Mg } Na_2SO_4}{4 \times 58.4425 \text{ Mg NaCl}}\right) = \mathbf{905 \text{ Mg}}$$

$$? \text{ Mg } Na_2SO_4 = 150 \text{ Mg } H_2O\left(\frac{2 \times 142.043 \text{ Mg } Na_2SO_4}{2 \times 18.0153 \text{ Mg } H_2O}\right) = 1.18 \times 10^3 \text{ Mg}$$

c. If 868 Mg Na_2SO_4 is formed in the reaction of 745 Mg of sodium chloride with 150 Mg of H_2O and excess SO_2 and O_2, what is the percent yield?

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{868 \text{ Mg}}{905 \text{ Mg}} \times 100 = \mathbf{95.9\% \text{ yield}}$$

d. What volume of 2.0 M Na_2SO_4 can be formed from the reaction of 745 Mg of sodium chloride with excess SO_2, H_2O, and O_2?

$$? \text{ Mg } Na_2SO_4 = 745 \text{ Mg NaCl}\left(\frac{10^6 \text{ g}}{1 \text{ Mg}}\right)\left(\frac{1 \text{ mol NaCl}}{58.4425 \text{ g NaCl}}\right)\left(\frac{2 \text{ mol } Na_2SO_4}{4 \text{ mol NaCl}}\right)\left(\frac{1 \text{ L } Na_2SO_4 \text{ soln}}{2.0 \text{ mol } Na_2SO_4}\right)$$

= 3.2×10^6 L Na_2SO_4 solution

Key Ideas Answers

4. Under typical conditions, the average distance between gas particles is about **ten** times their diameter.
6. For a gas at room temperature and pressure, the gas particles themselves occupy about **0.1%** of the total volume. The other **99.9%** of the total volume is empty space (whereas in liquids, about **70%** of the volume is occupied by particles).
8. The particles in a gas are in **rapid** and **continuous** motion.
10. The particles in a gas are constantly colliding with the walls of the container and with each other. Because of these collisions, the gas particles are constantly changing their **direction of motion** and their **velocity**.
12. The particles of an ideal gas are assumed to be **point-masses**, that is, particles that have a mass but occupy no volume.

14. The ideal gas model is used to predict changes in four related gas properties: **pressure**, **volume**, **number of particles**, and **temperature**.
16. Although gas temperatures are often measured with thermometers that report temperatures in **degrees Celsius, °C,** scientists generally use **Kelvin** temperatures for calculations.
18. The accepted SI unit for gas pressure is the **pascal, Pa**.
20. The observation that the pressure of an ideal gas is inversely proportional to the volume it occupies if the **moles of gas** and the temperature are constant is a statement of Boyle's Law.
22. The pressure of an ideal gas is directly proportional to the **Kelvin temperature** of the gas if the volume and the number of gas particles are constant. This relationship is sometimes called Gay-Lussac's Law.
24. For an ideal gas, volume and temperature described in kelvins are **directly** proportional if the number of gas particles and pressure are constant. This is a statement of Charles' Law.
26. If the **temperature and volume** of an ideal gas are held constant, the number of gas particles in a container and the gas pressure are directly proportional.
28. It is always a good idea to include the units in a solved equation as well as the numbers. If the units cancel to yield a reasonable unit for the unknown property, you can feel confident that you picked the **correct equation**, that you did the **algebra** correctly to solve for your unknown, and that you have made the **necessary unit conversions**.
30. There are three different ways to convert between a measurable property and moles in equation stoichiometry problems. For pure liquids and solids, we can convert between mass and moles, using the **molar mass** as a conversion factor. For gases, we can convert between volume of gas and moles using the methods described above. For solutions, **molarity** provides a conversion factor that enables us to convert between moles of solute and volume of solution.
32. Assuming ideal gas character, the partial pressure of any gas in a mixture is the pressure that the gas would exert if it were **alone** in the container.

Problems Key

Section 11.1 Gases and Their Properties

34. For a gas under typical conditions, what approximate percentage of the volume is occupied by the gas particles themselves? *(Obj 2)*

    0.1%

36. Why is it harder to walk through water than to walk through air?

    When we walk through air, we push the air particles out of the way as we move. Because the particles in a liquid occupy about 70% of the space that contains the liquid (as opposed to 0.1% of the space occupied by gas particles), there are a lot more particles to push out of the way as you move through water.

38. What are the key assumptions that distinguish an ideal gas from a real gas? *(Obj 4)*
    - *The particles are assumed to be point-masses, that is, particles that have a mass but occupy no volume.*
    - *There are no attractive or repulsive forces between the particles in an ideal gas.*

40. A TV weather person predicts a storm for the next day and refers to the dropping barometric pressure to support the prediction. What causes the pressure in the air?

The particles in the air (N_2, O_2, Xe, CO_2, and others) are constantly moving and constantly colliding with everything surrounded by the air. Each of these collisions exerts a tiny force against the object with which they collide. The total force of these collisions per unit area is the atmospheric pressure.

42. The pressure at the center of the earth is 4×10^{11} pascals, Pa. What is this pressure in kPa, atm, mmHg, and torr? *(Obj 8)*

$$? \text{ kPa} = 4\times10^{11} \text{ Pa}\left(\frac{1\text{ kPa}}{10^3\text{ Pa}}\right) = \mathbf{4 \times 10^8\ kPa}$$

$$? \text{ atm} = 4\times10^{11} \text{ Pa}\left(\frac{1\text{ atm}}{101325\text{ Pa}}\right) = \mathbf{4 \times 10^6\ atm}$$

$$? \text{ mmHg} = 4\times10^{11} \text{ Pa}\left(\frac{1\text{ atm}}{101325\text{ Pa}}\right)\left(\frac{760\text{ mmHg}}{1\text{ atm}}\right) = \mathbf{3 \times 10^9\ mmHg = 3 \times 10^9\ torr}$$

44. What does *inversely proportional* mean?

X and Y are inversely proportional if a decrease in X leads to a proportional increase in Y or an increase in X leads to a proportional decrease in Y. For example, volume of gas and its pressure are inversely proportional if the temperature and the number of gas particles are constant. If the volume is decreased to one-half its original value, the pressure of the gas will double. If the volume is doubled, the pressure decreases to one-half its original value. The following expression summarizes this inverse relationship:

$$P \propto \frac{1}{V} \quad \text{if n and T are constant}$$

47. Ammonia, NH_3, is a gas that can be dissolved in water and used as a cleaner. When an ammonia solution is used to clean the wax off a floor, ammonia gas escapes from the solution, mixes easily with the gas particles in the air, and spreads throughout the room. Very quickly, everyone in the room can smell the ammonia gas. Explain why gaseous ammonia mixes so easily with air.

In gases, there is plenty of empty space between the particles and essentially no attractions between them, so there is nothing to stop gases, such as ammonia and the gases in air, from mixing readily and thoroughly.

49. Explain why air moves in and out of our lungs as we breathe. *(Obj 12)*

When the muscles of your diaphragm contract and your chest expands, the volume of your lungs increases. This leads to a decrease in the number of particles per unit volume inside the lungs, which leaves fewer particles near any given area of the inner surface of the lungs. There are then fewer collisions per second per unit area of lungs and a decrease in force per unit area or gas pressure. During quiet, normal breathing, this increase in volume decreases the pressure in the lungs to about 0.4 kilopascals lower than the atmospheric pressure. The larger volume causes air to move into the lungs faster than it moves out, bringing in fresh oxygen. When the muscles relax, the lungs return to their original volume, and the decrease in volume causes the pressure in the lungs to increase to about 0.4 kilopascals above atmospheric pressure. Air now goes out of the lungs faster than it comes in. See Figure 11.9 of the textbook.

51. With reference to the relationships between properties of gases, answer the following questions.
    a. In a common classroom demonstration, water is heated in a 1-gallon can, covered with a tight lid, and allowed to cool. As the can cools, it collapses. Why?

       As the can cools, the water vapor in the can condenses to liquid, leaving fewer moles of gas. Decreased number of gas particles and decreased temperature both lead to decreased gas pressure in the can. Because the external pressure pushing on the outside of the can is then greater than the internal pressure pushing outward, the can collapses.
    b. Dents in ping-pong balls can often be removed by placing the balls in hot water. Why?

       The increased temperature causes the internal pressure of the ping-pong ball to increase. This leads to the pressure pushing out on the shell of the ball to be greater than the external pressure pushing in on the shell. If the difference in pressure is enough, the dents are pushed out.

Section 11.2 Calculations Involving Ideal Gases

53. Neon gas produced for use in luminous tubes comes packaged in 1.0-L containers at a pressure of 1.00 atm and a temperature of 19 °C. How many moles of Ne do these cylinders hold? (Obj 15)

$$PV = nRT \qquad n = \frac{PV}{RT} = \frac{1.00\text{ atm}\left(1.0\text{ L}\right)}{\left(\dfrac{0.082058\text{ L}\bullet\text{atm}}{\text{K}\bullet\text{mol}}\right)292\text{ K}} \mathbf{= 0.042\ moles\ Ne}$$

55. A typical aerosol can is able to withstand 10-12 atm without exploding.
    a. If a 375-mL aerosol can contains 0.062 mole of gas, at what temperature would the gas pressure reach 12 atm of pressure? (Obj 15)

$$PV = nRT \qquad T = \frac{PV}{nR} = \frac{12\text{ atm}\left(375\text{ mL}\right)}{0.062\text{ mol}\left(\dfrac{0.082058\text{ L}\bullet\text{atm}}{\text{K}\bullet\text{mol}}\right)}\left(\frac{1\text{ L}}{10^3\text{ mL}}\right)$$

$$\mathbf{= 8.8 \times 10^2\ K\ \ or\ \ 6.1 \times 10^2\ {}^\circ C}$$

    b. Aerosol cans usually contain liquids as well as gas. If the 375-mL can described in Part (a) contained liquid along with the 0.062 moles of gas, why would it explode at a lower temperature than if it contained only the gas? (*Hint:* What happens to a liquid when it is heated?) (Obj 15)

       As the temperature is increased, the liquid would evaporate more rapidly, increasing the amount of gas in the container and increasing the pressure.

57. Bromomethane, CH_3Br, commonly called methyl bromide, is used as a soil fumigant, but because it is a threat to the ozone layer, its use is being phased out. This colorless gas is toxic by ingestion, inhalation, or skin absorption, so the American Conference of Government Industrial Hygienists has assigned it a threshold limit value, or TLV, a maximum concentration to which a worker may be repeatedly exposed day after day without experiencing adverse effects. (TLV standards are meant to serve as guides in control of health hazards, not to provide definitive dividing lines between safe and dangerous concentrations.) Bromomethane reaches its TLV when 0.028 mole escapes into a room that has a volume of 180 m^3 and a temperature of 18 °C. What is the pressure in atmospheres of the gas under those conditions? (There are 10^3 L per cubic meter.) *(Obj 15)*

$$PV = nRT \qquad P = \frac{nRT}{V} = \frac{0.028\,\text{mol}\left(\dfrac{0.082058\,\text{L}\bullet\text{atm}}{\text{K}\bullet\text{mol}}\right)291\,\text{K}}{180\,\text{m}^3}\left(\frac{1\,\text{m}^3}{10^3\,\text{L}}\right) = \mathbf{3.7\times10^{-6}\ atm}$$

59. Gaseous chlorine, Cl_2, which is used in water purification, is a dense, greenish-yellow substance with an irritating odor. It is formed from the electrolysis of molten sodium chloride. Chlorine reaches its TLV (defined in Problem 57) in a laboratory when 0.017 mole of Cl_2 is released, yielding a pressure of 7.5×10^{-4} mmHg at a temperature of 19 °C. What is the volume, in liters, of the room? *(Obj 15)*

$$PV = nRT \qquad V = \frac{nRT}{P} = \frac{0.017\,\text{mol}\left(\dfrac{0.082058\,\text{L}\bullet\text{atm}}{\text{K}\bullet\text{mol}}\right)292\,\text{K}}{7.5\times10^{-4}\,\text{mmHg}}\left(\frac{760\,\text{mmHg}}{1\,\text{atm}}\right) = \mathbf{4.1\times10^{5}\ L}$$

61. Hydrogen sulfide, H_2S, which is used to precipitate sulfides of metals, is a colorless gas that smells like rotten eggs. It can be detected by the nose at about 0.6 mg/m^3. To reach this level in a 295-m^3 room, 0.177 g of H_2S must escape into the room. What is the pressure in pascals of 0.177 g of H_2S that escapes into a 295-m^3 laboratory room at 291 K? *(Obj 16)*

$$PV = \frac{g}{M}RT \qquad P = \frac{gRT}{MV} = \frac{0.177\,\text{g}\left(\dfrac{8.3145\,\text{L}\bullet\text{kPa}}{\text{K}\bullet\text{mol}}\right)291\,\text{K}}{\dfrac{34.082\,\text{g}}{\text{mol}}\,295\,\text{m}^3}\left(\frac{1\,\text{m}^3}{10^3\,\text{L}}\right)\left(\frac{10^3\,\text{Pa}}{1\,\text{kPa}}\right) = \mathbf{0.0426\ Pa}$$

63. Hydrogen chloride, HCl, is used to make vinyl chloride, which is then used to make polyvinyl chloride (PVC) plastic. Hydrogen chloride can be purchased in liquefied form in cylinders that contain 8 lb of HCl. What volume, in liters, will 8.0 lb of HCl occupy if the compound becomes a gas at 1.0 atm and 85 °C? *(Obj 16)*

$$PV = \frac{g}{M}RT \qquad V = \frac{gRT}{MP} = \frac{8.0\,\text{lb}\left(\dfrac{0.082058\,\text{L}\bullet\text{atm}}{\text{K}\bullet\text{mol}}\right)358\,\text{K}}{\dfrac{36.4606\,\text{g}}{1\,\text{mol}}\,1.0\,\text{atm}}\left(\frac{453.6\,\text{g}}{1\,\text{lb}}\right) = \mathbf{2.9\times10^{3}\ L}$$

65. The bulbs used for fluorescent lights have a mercury gas pressure of 1.07 Pa at 40 °C. How many milligrams of liquid mercury must evaporate at 40 °C to yield this pressure in a 1.39-L fluorescent bulb? *(Obj 16)*

$$PV = \frac{g}{M}RT \qquad g = \frac{PVM}{RT} = \frac{1.07\ \text{Pa}\,(1.39\ \text{L})\,\dfrac{200.59\ \text{g}}{1\ \text{mol}}}{\left(\dfrac{8.3145\ \text{L}\bullet\text{kPa}}{\text{K}\bullet\text{mol}}\right)313\ \text{K}}\left(\frac{1\ \text{kPa}}{10^3\ \text{Pa}}\right)\left(\frac{10^3\ \text{mg}}{1\ \text{g}}\right) = \mathbf{0.115\ mg}$$

67. Flashtubes are light bulbs that produce a high-intensity flash of light that is very short in duration. They are used in photography and for airport approach lighting. The xenon gas, Xe, in flashtubes has large atoms that can quickly carry the heat away from the filament to stop the radiation of light. At what temperature does 1.80 g of Xe in a 0.625-L flashtube have a pressure of 75.0 kPa? *(Obj 16)*

$$PV = \frac{g}{M}RT \qquad T = \frac{PVM}{gR} = \frac{75.0\ \text{kPa}\,(0.625\ \text{L})\,\dfrac{131.29\ \text{g}}{1\ \text{mol}}}{1.80\ \text{g}\left(\dfrac{8.3145\ \text{L}\bullet\text{kPa}}{\text{K}\bullet\text{mol}}\right)} = \mathbf{411\ K\ or\ 138\ °C}$$

69. When 0.690 g of an unknown gas is held in an otherwise-empty 285-mL container, the pressure is 756.8 mmHg at 19 °C. What is the molecular mass of the gas? *(Obj 16)*

$$PV = \frac{g}{M}RT$$

$$M = \frac{gRT}{PV} = \frac{0.690\ \text{g}\left(\dfrac{0.082058\ \text{L}\bullet\text{atm}}{\text{K}\bullet\text{mol}}\right)292\ \text{K}}{756.8\ \text{mmHg}\,(285\ \text{mL})}\left(\frac{10^3\ \text{mL}}{1\ \text{L}}\right)\left(\frac{760\ \text{mmHg}}{1\ \text{atm}}\right) = \mathbf{58.3\ g\,/\,mol}$$

71. Consider two helium tanks used to fill children's balloons. The tanks have equal volume and are at the same temperature, but tank A is new and therefore has a higher pressure of helium than tank B, which has been used to fill many balloons. Which tank holds gas with a greater density, tank A or tank B? Support your answer.

Tank A – *The following shows that density (g/V) is proportional to pressure when temperature is constant. Higher pressure means higher density at a constant temperature.*

$$PV = \frac{g}{M}RT \qquad \frac{g}{V} = P\left(\frac{M}{RT}\right)$$

73. Butadiene, $CH_2CHCHCH_2$, a suspected carcinogen, is used to make plastics such as styrene-butadiene rubber (SBR). Because it is highly flammable and forms explosive peroxides when exposed to air, it is often shipped in insulated containers at about 2 °C. What is the density of butadiene when its pressure is 102 kPa and its temperature is 2 °C? *(Obj 17)*

$$PV = \frac{g}{M}RT \qquad \frac{g}{V} = \frac{PM}{RT} = \frac{102\ \text{kPa}\left(\dfrac{54.092\ \text{g}}{1\ \text{mol}}\right)}{\left(\dfrac{8.3145\ \text{L}\bullet\text{kPa}}{\text{K}\bullet\text{mol}}\right)275\ \text{K}} = \mathbf{2.41\ g/L}$$

75. In order to draw air into your lungs, your diaphragm and other muscles contract, increasing the lung volume. This lowers the air pressure of the lungs to below atmospheric pressure, and air flows in. When your diaphragm and other muscles relax, the volume of the lungs decreases, and air is forced out. *(Obj 28)*

   a. If the total volume of your lungs at rest is 6.00 L and the initial pressure is 759 mmHg, what is the new pressure if the lung volume is increased to 6.02 L?

   $$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} \quad \text{Assuming constant moles of gas and temperature, } P_1V_1 = P_2V_2$$

   $$P_2 = P_1\left(\frac{V_1}{V_2}\right) = 759\text{ mmHg}\left(\frac{6.00\text{ L}}{6.02\text{ L}}\right) = \mathbf{756\ mmHg}$$

   b. If the total volume of your lungs at rest is 6.00 L and the initial pressure is 759 mmHg, at what volume will the pressure be 763 mmHg?

   $$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} \quad \text{Assuming constant moles of gas and temperature, } P_1V_1 = P_2V_2$$

   $$V_2 = V_1\left(\frac{P_1}{P_2}\right) = 6.00\text{ L}\left(\frac{759\text{ mmHg}}{763\text{ mmHg}}\right) = \mathbf{5.97\ L}$$

77. A scuba diver has 4.5 L of air in his lungs 66 ft below the ocean surface, where the pressure is 3.0 atm. What would the volume of this gas be at the surface, where the pressure is 1.0 atm? If the diver's lungs can hold no more than 7.0 L without rupturing, is he in trouble if he doesn't exhale as he rises to the surface? *(Obj 18)*

   $$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} \qquad \text{Assuming constant moles of gas and temperature, } P_1V_1 = P_2V_2$$

   $$V_2 = V_1\left(\frac{P_1}{P_2}\right) = 4.5\text{ L}\left(\frac{3.0\text{ atm}}{1.0\text{ atm}}\right) = \mathbf{13.5\ L}$$

   The diver had better exhale as he goes up.

78. Picture a gas in the apparatus shown in Figure 11.2 of the textbook. If the volume and temperature remain constant, how would you change the number of gas particles to increase the pressure of the gas?

   Increased number of gas particles leads to increased pressure when temperature and volume are constant.

80. A balloon containing 0.62 mole of gas has a volume of 15 L. Assuming constant temperature and pressure, how many moles of gas does the balloon contain when enough gas leaks out to decrease the volume to 11 L? *(Obj 18)*

   $$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} \qquad \text{Assuming constant pressure and temperature, } \frac{V_1}{n_1} = \frac{V_2}{n_2}$$

   $$n_2 = n_1\left(\frac{V_2}{V_1}\right) = 0.62\text{ mol}\left(\frac{11\text{ L}}{15\text{ L}}\right) = \mathbf{0.45\ mol}$$

82. Consider a weather balloon with a volume on the ground of 113 m^3 at 101 kPa and 14 °C. *(Obj 18)*

a. If the balloon rises to a height where the temperature is –50 °C and the pressure is 2.35 kPa, what will its volume be? (In 1958, a balloon reached 101,500 ft, or 30.85 km, the height at which the temperature is about –50 °C.)

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} \qquad \text{Assuming constant moles of gas, } \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = V_1\left(\frac{T_2}{T_1}\right)\left(\frac{P_1}{P_2}\right) = 113\ \text{m}^3\left(\frac{223\ \text{K}}{287\ \text{K}}\right)\left(\frac{101\ \text{kPa}}{2.35\ \text{kPa}}\right)\ \mathbf{= 3.77 \times 10^3\ m^3}$$

b. When the balloon rises to 10 km, where the temperature is –40 °C, its volume increases to 129 m^3. If the pressures inside and outside the balloon are equal, what is the atmospheric pressure at this height?

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} \qquad \text{Assuming constant moles of gas, } \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$P_2 = P_1\left(\frac{T_2}{T_1}\right)\left(\frac{V_1}{V_2}\right) = 101\ \text{kPa}\left(\frac{233\ \text{K}}{287\ \text{K}}\right)\left(\frac{113\ \text{m}^3}{129\ \text{m}^3}\right)\ \mathbf{= 71.8\ kPa}$$

c. If the balloon has a volume of 206 m^3 at 25 km, where the pressure is 45 kPa, what is the temperature at this height?

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} \qquad \text{Assuming constant moles of gas, } \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$T_2 = T_1\left(\frac{P_2}{P_1}\right)\left(\frac{V_2}{V_1}\right) = 287\ \text{K}\left(\frac{45\ \text{kPa}}{101\ \text{kPa}}\right)\left(\frac{206\ \text{m}^3}{113\ \text{m}^3}\right)\ \mathbf{= 233\ K}$$

Section 11.3 Equation Stoichiometry and Ideal Gases

85. Chlorine gas is used in the production of many other chemicals, including carbon tetrachloride, polyvinyl chloride plastic, and hydrochloric acid. It is produced from the electrolysis of molten sodium chloride. What minimum mass of sodium chloride, in megagrams, is necessary to make 2.7×10^5 L of chlorine gas at standard temperature and pressure, STP? *(Obj 18)*

$$2\text{NaCl}(l) \xrightarrow{\text{Electrolysis}} 2\text{Na}(l) + \text{Cl}_2(g)$$

$$?\ \text{Mg NaCl} = 2.7 \times 10^5\ \text{L Cl}_2\left(\frac{1\ \text{mol Cl}_2}{22.414\ \text{L Cl}_2}\right)_{\text{STP}}\left(\frac{2\ \text{mol NaCl}}{1\ \text{mol Cl}_2}\right)\left(\frac{58.4425\ \text{g NaCl}}{1\ \text{mol NaCl}}\right)\left(\frac{1\ \text{Mg}}{10^6\ \text{g}}\right)$$

$$\mathbf{= 1.4\ Mg\ NaCl}$$

87. Air bags in cars inflate when sodium azide, NaN_3, decomposes to generate nitrogen gas. How many grams of NaN_3 must react to generate 112 L N_2 at 121 kPa and 305 K? *(Obj 18)*

$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$

$$? \text{ g NaN}_3 = 112 \text{ L N}_2 \left(\frac{\text{K} \bullet \text{mol}}{8.3145 \text{ L} \bullet \text{kPa}}\right)\left(\frac{121 \text{ kPa}}{305 \text{ K}}\right)\left(\frac{2 \text{ mol NaN}_3}{3 \text{ mol N}_2}\right)\left(\frac{65.0099 \text{ g NaN}_3}{1 \text{ mol NaN}_3}\right)$$

= 232 g NaN_3

89. The hydrogen chloride gas used to make polyvinyl chloride (PVC) plastic is made by reacting hydrogen gas with chlorine gas. What volume of HCl(g) at 105 kPa and 296 K can be formed from 1150 m^3 of $H_2(g)$ at STP? *(Obj 18)*

$H_2(g) + Cl_2(g) \rightarrow 2 HCl(g)$

$$? \text{ m}^3 \text{ HCl} = 1150 \text{ m}^3 \text{ H}_2 \left(\frac{10^3 \text{ L}}{1 \text{ m}^3}\right)\left(\frac{1 \text{ mol H}_2}{22.414 \text{ L H}_2}\right)_{\text{STP}}\left(\frac{2 \text{ mol HCl}}{1 \text{ mol H}_2}\right)$$

$$\left(\frac{8.3145 \text{ L} \bullet \text{kPa}}{\text{K} \bullet \text{mol}}\right)\left(\frac{296 \text{ K}}{105 \text{ kPa}}\right)\left(\frac{1 \text{ m}^3}{10^3 \text{ L}}\right) = \mathbf{2.41 \times 10^3 \ m^3 \ HCl}$$

91. Ammonia and carbon dioxide combine to produce urea, NH_2CONH_2, and water through the following reaction. About 90% of the urea is used to make fertilizers, but some is also used to make animal feed, plastics, and pharmaceuticals. To get the optimum yield of urea, the ratio of NH_3 to CO_2 is carefully regulated at a 3:1 molar ratio of NH_3 to CO_2.

$2NH_3(g) + CO_2(g) \rightarrow NH_2CONH_2(s) + H_2O(l)$

a. What volume of NH_3 at STP must be combined with 1.4×10^4 L of CO_2 at STP to yield the 3:1 molar ratio? *(Obj 18)*

$$? \text{ L NH}_3 = 1.4 \times 10^4 \text{ CO}_2 \left(\frac{1 \text{ mol CO}_2}{22.414 \text{ L CO}_2}\right)_{\text{STP}}\left(\frac{3 \text{ mol NH}_3}{1 \text{ mol CO}_2}\right)\left(\frac{22.414 \text{ L NH}_3}{1 \text{ mol NH}_3}\right)_{\text{STP}}$$

= 4.2×10^4 L NH_3

b. The concentration of urea in a typical liquid fertilizer is about 1.0 M NH_2CONH_2. What minimum volume of CO_2 at 190 °C and 34.0 atm is needed to make 3.5×10^4 L of 1.0 M NH_2CONH_2? *(Obj 18)*

$$? \text{ L CO}_2 = 3.5 \times 10^4 \text{ L NH}_2\text{CONH}_2 \text{ soln}\left(\frac{1.0 \text{ mol NH}_2\text{CONH}_2}{1 \text{ L NH}_2\text{CONH}_2 \text{ soln}}\right)\left(\frac{1 \text{ mol CO}_2}{1 \text{ mol NH}_2\text{CONH}_2}\right)$$

$$\left(\frac{0.082058 \text{ L} \bullet \text{atm}}{\text{K} \bullet \text{mol}}\right)\left(\frac{463 \text{ K}}{34.0 \text{ atm}}\right) = \mathbf{3.9 \times 10^4 \ L \ CO_2}$$

93. Some chlorofluorocarbons, CFCs, are formed from carbon tetrachloride, CCl_4, which can be made from the methane in natural gas, as shown below. What minimum volume in cubic meters of methane gas at STP must be added to react completely 1.9×10^6 L of $Cl_2(g)$ at STP? *(Obj 18)*

$CH_4(g) + 4Cl_2(g) \rightarrow CCl_4(l) + 4HCl(g)$

$$? \text{ m}^3 \text{ CH}_4 = 1.9 \times 10^6 \text{ L Cl}_2 \left(\frac{1 \text{ mol Cl}_2}{22.414 \text{ L Cl}_2}\right)_{\text{STP}}\left(\frac{1 \text{ mol CH}_4}{4 \text{ mol Cl}_2}\right)\left(\frac{22.414 \text{ L CH}_4}{1 \text{ mol CH}_4}\right)_{\text{STP}}\left(\frac{1 \text{ m}^3}{10^3 \text{ L}}\right)$$

= 475 m^3 CH_4

95. Ethylene oxide, C_2H_4O, is a cyclic compound; its two carbon atoms and its oxygen atom form a three-member ring. The small ring is somewhat unstable, so ethylene oxide is a very reactive substance. It is useful as a rocket propellant, to sterilize medical plastic tubing, and to make polyesters and antifreeze. It is formed from the reaction of ethylene, CH_2CH_2, with oxygen at 270-290 °C and 8-20 atm in the presence of a silver catalyst.

$$2CH_2CH_2(g) + O_2(g) \rightarrow 2C_2H_4O(g)$$

a. What minimum mass, in megagrams, of liquefied ethylene, CH_2CH_2, is necessary to form 2.0×10^2 m^3 of ethylene oxide, C_2H_4O, at 107 kPa and 298 K? *(Obj 18)*

$$? \text{ Mg } CH_2CH_2 = 2.0\times10^2 \text{ m}^3 \text{ } C_2H_4O\left(\frac{10^3 \text{ L}}{1 \text{ m}^3}\right)\left(\frac{\text{K}\bullet\text{mol}}{8.3145 \text{ L}\bullet\text{kPa}}\right)\left(\frac{107 \text{ kPa}}{298 \text{ K}}\right)$$
$$\left(\frac{2 \text{ mol } CH_2CH_2}{2 \text{ mol } C_2H_4O}\right)\left(\frac{28.054 \text{ g } CH_2CH_2}{1 \text{ mol } CH_2CH_2}\right)\left(\frac{1 \text{ Mg}}{10^6 \text{ g}}\right) = \mathbf{0.24\ Mg\ CH_2CH_2}$$

b. Ethylene can be purchased in cylinders containing 30 lb of liquefied ethylene. Assuming complete conversion of ethylene into ethylene oxide, determine how many of these cylinders would be necessary to make 2.0×10^2 m^3 of ethylene oxide at 107 kPa and 298 K? *(Obj 18)*

$$? \text{ cylinders} = 0.24 \text{ Mg } CH_2CH_2\left(\frac{10^6 \text{ g}}{1 \text{ Mg}}\right)\left(\frac{1 \text{ lb}}{453.6 \text{ g}}\right)\left(\frac{1 \text{ cylinder}}{30 \text{ lb } CH_2CH_2}\right) = \mathbf{18\ cylinders}$$

c. How many liters of ethylene oxide gas at 2.04 atm and 67 °C can be made from 1.4×10^4 L of CH_2CH_2 gas at 21 °C and 1.05 atm? *(Obj 18)*

$$? \text{ L } C_2H_4O = 1.4\times10^4 \text{ L } CH_2CH_2\left(\frac{\text{K}\bullet\text{mol}}{0.082058 \text{ L}\bullet\text{atm}}\right)\left(\frac{1.05 \text{ atm}}{294 \text{ K}}\right)\left(\frac{2 \text{ mol } C_2H_4O}{2 \text{ mol } CH_2CH_2}\right)$$
$$\left(\frac{0.082058 \text{ L}\bullet\text{atm}}{\text{K}\bullet\text{mol}}\right)\left(\frac{340 \text{ K}}{2.04 \text{ atm}}\right) = \mathbf{8.3\times10^3\ L\ C_2H_4O}$$

d. How many kilograms of liquefied ethylene oxide can be made from the reaction of 1.4×10^4 L of CH_2CH_2 gas at 1.05 atm and 21 °C and 8.9×10^3 L of oxygen gas at 0.998 atm and 19 °C? *(Obj 18)*

$$? \text{ kg } C_2H_4O = 1.4\times10^4 \text{ L } CH_2CH_2\left(\frac{\text{K}\bullet\text{mol}}{0.082058 \text{ L}\bullet\text{atm}}\right)\left(\frac{1.05 \text{ atm}}{294 \text{ K}}\right)\left(\frac{2 \text{ mol } C_2H_4O}{2 \text{ mol } CH_2CH_2}\right)$$
$$\left(\frac{44.053 \text{ g } C_2H_4O}{1 \text{ mol } C_2H_4O}\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right) = \mathbf{27\ kg\ C_2H_4O}$$

$$? \text{ kg } C_2H_4O = 8.9\times10^3 \text{ L } O_2\left(\frac{\text{K}\bullet\text{mol}}{0.082058 \text{ L}\bullet\text{atm}}\right)\left(\frac{0.998 \text{ atm}}{292 \text{ K}}\right)\left(\frac{2 \text{ mol } C_2H_4O}{1 \text{ mol } O_2}\right)$$
$$\left(\frac{44.053 \text{ g } C_2H_4O}{1 \text{ mol } C_2H_4O}\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right) = 33 \text{ kg } C_2H_4O$$

97. Rutile ore is about 95% titanium(IV) oxide, TiO_2. The TiO_2 is purified in the two reactions that follow. In the first reaction, what minimum volume, in cubic meters, of Cl_2 at STP is necessary to convert all of the TiO_2 in 1.7×10^4 metric tons of rutile ore into $TiCl_4$?
(Obj 18)

$$3TiO_2 + 4C + 6Cl_2 \xrightarrow{900\ °C} 3TiCl_4 + 2CO + 2CO_2$$

$$TiCl_4 + O_2 \xrightarrow{1200\text{-}1400\ °C} TiO_2 + 2Cl_2$$

$$?\ m^3\ Cl_2 = 1.7\times10^4\ t\ ore\left(\frac{10^3\ kg}{1\ t}\right)\left(\frac{10^3\ g}{1\ kg}\right)\left(\frac{95\ g\ TiO_2}{100\ g\ ore}\right)\left(\frac{1\ mol\ TiO_2}{79.866\ g\ TiO_2}\right)$$

$$\left(\frac{6\ mol\ Cl_2}{3\ mol TiO_2}\right)\left(\frac{22.414\ L\ Cl_2}{1\ mol\ Cl_2}\right)_{STP}\left(\frac{1\ m^3}{10^3\ L}\right) = \mathbf{9.1\times10^6\ m^3\ Cl_2}$$

Section 11.4 Dalton's Law of Partial Pressures

100. A typical 100-watt incandescent light bulb contains argon gas and nitrogen gas.
(Objs 21 & 22)

a. If enough argon and nitrogen are added to this light bulb to yield a partial pressure of Ar of 98.5 kPa and a partial pressure of N_2 of 10.9 kPa, what is the total gas pressure in the bulb?

$$P_{total} = P_{Ar} + P_{N_2} = 98.5\ kPa + 10.9\ kPa = \mathbf{109.4\ kPa}$$

b. If a 125-mL light bulb contains 5.99×10^{-4} mole of N_2 and 5.39×10^{-3} mole of Ar, what is the pressure in kPa in the bulb at a typical operating temperature of 119 °C?

$$P_{total} = \left(\sum n\right)\frac{RT}{V} = \left(5.99\times10^{-4}\ mol + 5.39\times10^{-4}\ mol\right)\frac{\frac{8.3145\ L \bullet kPa}{K \bullet mol}(392\ K)}{125\ mL}\left(\frac{10^3\ mL}{1\ L}\right)$$

$$= \mathbf{29.7\ kPa}$$

102. If a 135-mL luminous tube contains 1.2 mg of Ne and 1.2 mg of Ar, at what temperature will the total gas pressure be 13 mmHg?

$$?\ mol\ Ne = 1.2\ mg\ Ne\left(\frac{1\ g}{10^3\ mg}\right)\left(\frac{1\ mol\ Ne}{20.1797\ g\ Ne}\right) = 5.9 \times 10^{-5}\ mol\ Ne$$

$$?\ mol\ Ar = 1.2\ mg\ Ar\left(\frac{1\ g}{10^3\ mg}\right)\left(\frac{1\ mol\ Ar}{39.948\ g\ Ar}\right) = 3.0 \times 10^{-5}\ mol\ Ar$$

$$P_{total} = \left(\sum n\right)\frac{RT}{V}$$

$$T = \frac{P_{total}V}{\left(\sum n\right)R} = \frac{13\ mmHg\ (135\ mL)}{\left(5.9\times10^{-5}\ mol + 3.0\times10^{-5}\ mol\right)\frac{0.082058\ L \bullet atm}{K \bullet mol}}\left(\frac{1\ atm}{760\ mmHg}\right)\left(\frac{1\ L}{10^3\ mL}\right)$$

$$= \mathbf{316\ K\ or\ 43\ °C}$$

104. Natural gas is a mixture of gaseous hydrocarbons and other gases in small quantities. The percentage of each component varies depending on the source. Consider a 21.2-m^3 storage container that holds 11.8 kg of methane gas, CH_4, 2.3 kg of ethane gas, C_2H_6, 1.1 kg of propane gas, $C_3H_8(g)$, and an unknown amount of other gases.

   a. If the total pressure in the container is 1.00 atm at 21 °C, how many moles of gases other than CH_4, C_2H_6, and C_3H_8 are in the container?

$$P_{total} = \left(\sum n\right)\frac{RT}{V} \qquad \left(\sum n\right) = \frac{P_{total}V}{RT} = \frac{1.00\text{ atm}\left(21.2\text{ m}^3\right)}{\left(\dfrac{0.082058\text{ L}\bullet\text{atm}}{\text{K}\bullet\text{mol}}\right)294\text{ K}}\left(\frac{10^3\text{ L}}{1\text{ m}^3}\right) = 879\text{ mol}$$

$$?\text{ mol CH}_4 = 11.8\text{ kg CH}_4\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol CH}_4}{16.043\text{ g CH}_4}\right) = 736\text{ mol CH}_4$$

$$?\text{ mol C}_2\text{H}_6 = 2.3\text{ kg C}_2\text{H}_6\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol C}_2\text{H}_6}{30.070\text{ g C}_2\text{H}_6}\right) = 76\text{ mol C}_2\text{H}_6$$

$$?\text{ mol C}_3\text{H}_8 = 1.1\text{ kg C}_3\text{H}_8\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol C}_3\text{H}_8}{44.097\text{ g C}_3\text{H}_8}\right) = 25\text{ mol C}_3\text{H}_8$$

$$?\text{ mol other} = n_{total} - n_{CH_4} - n_{C_2H_6} - n_{C_3H_8} = 879\text{ mol} - 736\text{ mol} - 76\text{ mol} - 25\text{ mol}$$

$$= \mathbf{42\text{ mol other}}$$

   b. What percentage of the gas particles in the cylinder are methane molecules?

$$?\,\%\text{ CO}_2 = \frac{736\text{ mol CH}_4}{879\text{ mol total}}\times 100 = \mathbf{83.7\%}$$

Additional Problems

106. Although the temperature decreases with increasing altitude in the troposphere (the lowest portion of the earth's atmosphere), it increases again in the upper portion of the stratosphere. Thus, at about 50 km, the temperature is back to about 20 °C. Compare the following properties of the air along the California coastline at 20 °C and air at the same temperature but an altitude of 50 km. Assume that the percent composition of the air in both places is essentially the same.

   a. Do the particles in the air along the California coastline and the air at 50 km have the same average kinetic energy? If not, which has particles with the higher average kinetic energy? Explain your answer.

      Yes. The average kinetic energy is dependent on the temperature of the particles, so the average kinetic energy is the same when the temperature is the same.

   b. Do the particles in the air along the California coastline and the air at 50 km have the same velocity? If not, which has particles with the higher average velocity? Explain your answer.

      Yes. If the composition of the air is the same, the average mass of the particles is the same. Because the average kinetic energy of the particles is the same for the same temperature, the average velocity must also be the same.

$$KE_{average} = \tfrac{1}{2}\, m u^2_{average}$$

c. Does the air along the California coastline and the air at 50 km have the same average distance between the particles? If not, which has the greater average distance between the particles? Explain your answer.

No. Because there are fewer particles per unit volume at 50 km, the average distance between the particles is greater than that at sea level.

d. Do the particles in the air along the California coastline and the air at 50 km have the same average frequency of particle collisions? If not, which has the greater frequency of between particle collisions? Explain your answer.

No. Because the average velocity of the particles is the same and the average distance between the particles at 50 km is longer, the average distance between the collisions at 50 km is greater than that at sea level. Thus the frequency of collisions is greater at sea level.

e. Does the air along the California coastline and the air at 50 km have the same density? If not, which has the higher density? Explain your answer.

No, because there are fewer particles per unit volume, the density of the air at 50 km is much less than at sea level. Thus the sea level air has the higher density.

f. Does the air along the California coastline and the air at 50 km have the same gas pressure? If not, which has the higher gas pressure? Explain your answer.

No. Because the particles in the air in the two locations would have the same average mass and the same average velocity, they would collide with the walls of a container with the same force per collision. Because there are more gas particles per unit volume at sea level than at 50 km, there would be more collisions with the walls of a container that holds a sample of the air along the California coastline. This would lead to a greater force pushing on the walls of the container and a greater gas pressure for the gas at sea level.

108. With reference to the relationships between properties of gases, answer the following questions.

a. The pressure in the tires of a car sitting in a garage is higher at noon than at midnight. Why?

The higher temperature at noon causes the pressure to increase.

b. The pressure in a car's tires increases as it is driven from home to school. Why?

The interaction between the moving tires and the stationary road causes the particles in the tires to increase their velocity, so the temperature of the tires and ultimately of the gas in the tires goes up. The increase in the temperature of the gas in the tire increases the pressure of the gas.

110. When people inhale helium gas and talk, their voices sound strange, as if it had been recorded and played back at a faster rate. Why do you think that happens?

Helium is less dense than air, so it is easier for our vocal cords to vibrate. Because of this, they vibrate faster, making our voices sound strange.