Equation 11.5 (page 520) Density of gases (where \( d \) is the gas density in \( \text{g} \text{L}^{-1} \) is the molar mass of the gas).

\[
d = \frac{m}{V} = \frac{PM}{RT}
\]

Equation 11.6 (page 525) Dalton’s law of partial pressures. The total pressure of a gas mixture is the sum of the partial pressures of the component gases (\( P_s \)).

\[
P_{\text{total}} = P_1 + P_2 + P_3 + \ldots
\]

Equation 11.7 (page 525) The total pressure of a gas mixture is equal to the number of moles of gases multiplied by (\( RT/V \)).

\[
P_{\text{total}} = (n_{\text{total}})(\frac{RT}{V})
\]

Equation 11.8 (page 526) The pressure of a gas (\( A \)) in a mixture is the product of its mole fraction (\( X_A \)) and the total pressure of the mixture.

\[
P_A = X_A P_{\text{total}}
\]

Equation 11.9 (page 530) The rms speed (\( \sqrt{\frac{u^2}{M}} \)) depends on the molar mass of a gas (\( M \)) and its temperature (\( T \)).

\[
\sqrt{\frac{u^2}{M}} = \sqrt{\frac{3RT}{M}}
\]

Equation 11.10 (page 533) Graham’s law. The rate of effusion of a gas is inversely proportional to the square root of its molar mass.

\[
\frac{\text{Rate of effusion of gas 1}}{\text{Rate of effusion of gas 2}} = \sqrt{\frac{\text{molar mass of gas 2}}{\text{molar mass of gas 1}}}
\]

Equation 11.11 (page 535) The van der Waals equation, which relates pressure, volume, temperature, and amount of gas for a nonideal gas.

\[
(P + a \left( \frac{n}{V} \right)^2)(V - bn) = nRT
\]

- **Study Questions**
  - Interactive versions of these questions are assignable in OWL.
  - \( \Delta \) denotes challenging questions.
  - Blue-numbered questions have answers in Appendix R and fully worked solutions in the Student Solutions Manual.

**Practicing Skills**

(See Section 11.1 and Example 11.1.)

1. The pressure of a gas is 440 mm Hg. Express this pressure in units of (a) atmospheres, (b) bars, and (c) kilopascals.

2. The average barometric pressure at an altitude of 10 km is 210 mm Hg. Express this pressure in atmospheres, bars, and kilopascals.

3. Indicate which represents the higher pressure in each of the following pairs:
   (a) 534 mm Hg or 0.754 bar
   (b) 534 mm Hg or 650 kPa
   (c) 1.34 bar or 934 kPa

4. Put the following in order of increasing pressure:
   363 mm Hg, 363 kPa, 0.256 atm, and 0.523 bar.

**Boyle’s Law and Charles’s Law**

(See Section 11.2 and Examples 11.2 and 11.3.)

5. A sample of nitrogen gas has a pressure of 67.5 mm Hg in a 500-mL flask. What is the pressure of this gas sample when it is transferred to a 125-mL flask at the same temperature?
6. A sample of CO$_2$ gas has a pressure of 56.5 mm Hg in a 125-mL flask. The sample is transferred to a new flask, where it has a pressure of 62.3 mm Hg at the same temperature. What is the volume of the new flask?

7. You have 3.5 L of NO at a temperature of 22.0 °C. What volume would the NO occupy at 37 °C? (Assume the pressure is constant.)

8. A 5.0-mL sample of CO$_2$ gas is enclosed in a gas-tight syringe (Figure 11.3) at 22 °C. If the syringe is immersed in an ice bath (0 °C), what is the new gas volume, assuming that the pressure is held constant?

The General Gas Law
(See Section 11.2 and Example 11.4.)

9. You have 3.6 L of H$_2$ gas at 380 mm Hg and 25 °C. What is the pressure of this gas if it is transferred to a 5.0-L flask at 0 °C?

10. You have a sample of CO$_2$ in flask A with a volume of 25.0 mL. At 20.5 °C, the pressure of the gas is 43.5 mm Hg. To find the volume of another flask, B, you move the CO$_2$ to that flask and find that its pressure is now 94.3 mm Hg at 24.5 °C. What is the volume of flask B?

11. You have a sample of gas in a flask with a volume of 250 mL. At 25.5 °C, the pressure of the gas is 360 mm Hg. If you decrease the temperature to −5.0 °C, what is the gas pressure at the lower temperature?

12. A sample of gas occupies 135 mL at 22.5 °C; the pressure is 165 mm Hg. What is the pressure of the gas sample when it is placed in a 252-mL flask at a temperature of 0.0 °C?

13. One of the cylinders of an automobile engine has a volume of 400. cm$^3$. The engine takes in air at a pressure of 1.00 atm and a temperature of 15 °C and compresses the air to a volume of 50.0 cm$^3$ at 77 °C. What is the final pressure of the gas in the cylinder? (The ratio of before and after volumes—in this case, 400:50 or 8:1—is called the compression ratio.)

14. A helium-filled balloon of the type used in long-distance flying contains 420,000 ft$^3$ (1.2 × 10$^7$ L) of helium. Suppose you fill the balloon with helium on the ground, where the pressure is 737 mm Hg and the temperature is 16.0 °C. When the balloon ascends to a height of 2 miles, where the pressure is only 600. mm Hg and the temperature is −33 °C, what volume is occupied by the helium gas? Assume the pressure inside the balloon matches the external pressure. Comment on the result.

Avogadro’s Hypothesis
(See Section 11.2 and Example 11.5.)

15. Nitrogen monoxide reacts with oxygen to give nitrogen dioxide.

2 NO(g) + O$_2$(g) → 2 NO$_2$(g)

(a) You wish to react NO and O$_2$ in the correct stoichiometric ratio. The sample of NO has a volume of 150 mL. What volume of O$_2$ is required (at the same pressure and temperature)?

(b) What volume of NO$_2$ (at the same pressure and temperature) is formed in this reaction?

16. Ethane burns in air to give H$_2$O and CO$_2$.

2 C$_2$H$_6$(g) + 7 O$_2$(g) → 4 CO$_2$(g) + 6 H$_2$O(g)

What volume of O$_2$ (L) is required for complete reaction with 5.2 L of C$_2$H$_6$? What volume of H$_2$O vapor (L) is produced? Assume all gases are measured at the same temperature and pressure.

Ideal Gas Law
(See Section 11.3 and Example 11.6.)

17. A 1.25-g sample of CO$_2$ is contained in a 750-mL flask at 22.5 °C. What is the pressure of the gas?

18. A balloon holds 30.0 kg of helium. What is the volume of the balloon if its pressure is 1.20 atm and the temperature is 22 °C?

19. A flask is first evacuated so that it contains no gas at all. Then, 2.2 g of CO$_2$ is introduced into the flask. On warming to 22 °C, the gas exerts a pressure of 318 mm Hg. What is the volume of the flask?

20. A steel cylinder holds 1.50 g of ethanol, C$_2$H$_5$OH. What is the pressure of the ethanol vapor if the cylinder has a volume of 251 cm$^3$ and the temperature is 250 °C? (Assume all of the ethanol is in the vapor phase at this temperature.)

21. A balloon for long-distance flying contains 1.2 × 10$^7$ L of helium. If the helium pressure is 737 mm Hg at 25 °C, what mass of helium (in grams) does the balloon contain? (See Study Question 14.)

22. What mass of helium, in grams, is required to fill a 5.0-L balloon to a pressure of 1.1 atm at 25 °C?

Gas Density and Molar Mass
(See Section 11.3 and Examples 11.7 and 11.8.)

23. Forty miles above Earth’s surface, the temperature is 250 K, and the pressure is only 0.29 mm Hg. What is the density of air (in grams per liter) at this altitude? (Assume the molar mass of air is 28.96 g/mol.)

24. Diethyl ether, (C$_2$H$_5$)$_2$O, vaporizes easily at room temperature. If the vapor exerts a pressure of 233 mm Hg in a flask at 25 °C, what is the density of the vapor?

25. A gaseous organofluorine compound has a density of 0.355 g/L at 17 °C and 189 mm Hg. What is the molar mass of the compound?

26. Chloroform is a common liquid used in the laboratory. It vaporizes readily. If the pressure of chloroform vapor in a flask is 195 mm Hg at 25.0 °C and the density of the vapor is 1.25 g/L, what is the molar mass of chloroform?

27. A 1.007-g sample of an unknown gas exerts a pressure of 715 mm Hg in a 452-mL container at 23 °C. What is the molar mass of the gas?

28. A 0.0125-g sample of a gas with an empirical formula of CHF$_3$ is placed in a 165-mL flask. It has a pressure of 13.7 mm Hg at 22.5 °C. What is the molecular formula of the compound?
29. A new boron hydride, $\text{B}_2\text{H}_6$, has been isolated. To find its molar mass, you measure the pressure of the gas in a known volume at a known temperature. The following experimental data are collected:

- Mass of gas = 12.5 mg
- Pressure of gas = 24.8 mm Hg
- Temperature = 25 °C
- Volume of flask = 125 mL

Which formula corresponds to the calculated molar mass?

- (a) $\text{B}_2\text{H}_6$
- (b) $\text{B}_4\text{H}_{10}$
- (c) $\text{B}_6\text{H}_9$
- (d) $\text{B}_9\text{H}_{14}$
- (e) $\text{B}_{10}\text{H}_9$

30. Acetaldehyde is a common liquid compound that vaporizes readily. Determine the molar mass of acetaldehyde from the following data:

- Sample mass = 0.107 g
- Volume of gas = 125 mL
- Temperature = 0.0 °C
- Pressure = 331 mm Hg

**Gas Laws and Chemical Reactions**

*(See Section 11.4 and Examples 11.9 and 11.10.)*

31. Iron reacts with hydrochloric acid to produce iron(II) chloride and hydrogen gas:

$$\text{Fe}(s) + 2 \text{HCl(aq)} \rightarrow \text{FeCl}_2(\text{aq}) + \text{H}_2(\text{g})$$

The $\text{H}_2$ gas from the reaction of 2.2 g of iron with excess acid is collected in a 10.0-L flask at 25 °C. What is the pressure of the $\text{H}_2$ gas in this flask?

32. Silane, $\text{SiH}_4$, reacts with $\text{O}_2$ to give silicon dioxide and water:

$$\text{SiH}_4(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow \text{SiO}_2(\text{s}) + 2 \text{H}_2\text{O}(\ell)$$

A 5.20-L sample of $\text{SiH}_4$ gas at 356 mm Hg pressure and 25 °C is allowed to react with $\text{O}_2$ gas. What volume of $\text{O}_2$ gas, in liters, is required for complete reaction if the oxygen has a pressure of 425 mm Hg at 25 °C?

33. Sodium azide, the explosive compound in automobile air bags, decomposes according to the following equation:

$$2 \text{NaN}_3(s) \rightarrow 2 \text{Na}(s) + 3 \text{N}_2(g)$$

What mass of sodium azide is required to provide the nitrogen needed to inflate a 75.0-L bag to a pressure of 1.3 atm at 25 °C?

34. The hydrocarbon octane ($\text{C}_8\text{H}_{18}$) burns to give $\text{CO}_2$ and water vapor:

$$2 \text{C}_8\text{H}_{18}(g) + 25 \text{O}_2(g) \rightarrow 16 \text{CO}_2(g) + 18 \text{H}_2\text{O}(\ell)$$

If a 0.048-g sample of octane burns completely in $\text{O}_2$, what will be the pressure of water vapor in a 4.75-L flask at 30.0 °C? If the $\text{O}_2$ gas needed for complete combustion was contained in a 4.75-L flask at 22 °C, what would its pressure be?

35. Hydrazine reacts with $\text{O}_2$ according to the following equation:

$$\text{N}_2\text{H}_4(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + 2 \text{H}_2\text{O}(\ell)$$

Assume the $\text{O}_2$ needed for the reaction is in a 450-L tank at 25 °C. What must the oxygen pressure be in the tank to have enough oxygen to consume 1.00 kg of hydrazine completely?

36. A self-contained underwater breathing apparatus uses canisters containing potassium superoxide. The superoxide consumes the $\text{CO}_2$ exhaled by a person and replaces it with oxygen.

$$4 \text{KO}_2(s) + 2 \text{CO}_2(g) \rightarrow 2 \text{K}_2\text{CO}_3(s) + 3 \text{O}_2(g)$$

What mass of $\text{KO}_2$, in grams, is required to react with 8.90 L of $\text{CO}_2$ at 22.0 °C and 767 mm Hg?

**Gas Mixtures and Dalton’s Law**

*(See Section 11.5 and Example 11.11.)*

37. What is the total pressure in atmospheres of a gas mixture that contains 1.0 g of $\text{H}_2$ and 8.0 g of $\text{Ar}$ in a 3.0-L container at 27 °C? What are the partial pressures of the two gases?

38. A cylinder of compressed gas is labeled “Composition (mole %): 4.5% $\text{H}_2\text{O}$, 3.0% $\text{CO}_2$, balance $\text{N}_2$.” The pressure gauge attached to the cylinder reads 46 atm. Calculate the partial pressure of each gas, in atmospheres, in the cylinder.

39. A halothane–oxygen mixture ($\text{C}_2\text{HBrClF}_3 + \text{O}_2$) can be used as an anesthetic. A tank containing such a mixture has the following partial pressures: $P$ (halothane) = 170 mm Hg and $P$ (O$_2$) = 570 mm Hg.

(a) What is the ratio of the number of moles of halothane to the number of moles of $\text{O}_2$?

(b) If the tank contains 160 g of $\text{O}_2$, what mass of $\text{C}_2\text{HBrClF}_3$ is present?

40. A collapsed balloon is filled with He to a volume of 12.5 L at a pressure of 1.00 atm. Oxygen, $\text{O}_2$, is then added so that the final volume of the balloon is 26 L with a total pressure of 1.00 atm. The temperature, which remains constant throughout, is 21.5 °C.

(a) What mass of He does the balloon contain?

(b) What is the final partial pressure of He in the balloon?

(c) What is the partial pressure of $\text{O}_2$ in the balloon?

(d) What is the mole fraction of each gas?

**Kinetic-Molecular Theory**

*(See Section 11.6 and Example 11.12.)*

41. You have two flasks of equal volume. Flask A contains $\text{H}_2$ at 0 °C and 1 atm pressure. Flask B contains $\text{CO}_2$ gas at 25 °C and 2 atm pressure. Compare these two gases with respect to each of the following:

(a) average kinetic energy per molecule

(b) root mean square speed

(c) number of molecules

(d) mass of gas

42. Equal masses of gaseous $\text{N}_2$ and $\text{Ar}$ are placed in separate flasks of equal volume at the same temperature. Tell whether each of the following statements is true or false. Briefly explain your answer in each case.

(a) There are more molecules of $\text{N}_2$ present than atoms of $\text{Ar}$.

(b) The pressure is greater in the $\text{Ar}$ flask.

(c) The $\text{Ar}$ atoms have a greater rms speed than the $\text{N}_2$ molecules.

(d) The $\text{N}_2$ molecules collide more frequently with the walls of the flask than do the $\text{Ar}$ atoms.

43. If the rms speed of an oxygen molecule is $4.28 \times 10^4$ cm/s at 25 °C, what is the rms speed of a $\text{CO}_2$ molecule at the same temperature?
44. Calculate the rms speed for CO molecules at 25 °C. What is the ratio of this speed to that of Ar atoms at the same temperature?

45. Place the following gases in order of increasing rms speed at 25 °C: Ar, CH₄, N₂, CH₂F₂.

46. The reaction of SO₂ with Cl₂ gives dichlorine oxide, which is used to bleach wood pulp and to treat wastewater:

\[ \text{SO}_2(g) + 2 \text{Cl}_2(g) \rightarrow \text{OSCl}_2(g) + \text{Cl}_2\text{O}(g) \]

All of the compounds involved in the reaction are gases. List them in order of increasing rms speed.

**Diffusion and Effusion**

(See Section 11.7 and Example 11.13.)

47. In each pair of gases below, tell which will effuse faster:
   (a) CO₂ or F₂
   (b) O₂ or N₂
   (c) C₂H₄ or C₂H₆
   (d) two chlorofluorocarbons: CFC₁₃ or C₂ClF₄

48. Argon gas is 10 times denser than helium gas at the same temperature and pressure. Which gas is predicted to effuse faster? How much faster?

49. A gas whose molar mass you wish to know effuses through an opening at a rate one third as fast as that of helium gas. What is the molar mass of the unknown gas?

50. ▲ A sample of uranium fluoride is found to effuse at the rate of 17.7 mg/h. Under comparable conditions, gaseous I₂ effuses at the rate of 15.0 mg/h. What is the molar mass of the uranium fluoride? (Hint: Rates must be converted to units of moles per time.)

**Nonideal Gases**

(See Section 11.8)

51. In the text, it is stated that the pressure of 4.00 mol of Cl₂ in a 4.00-L tank at 100.0 °C should be 26.0 atm if calculated using the van der Waals equation. Verify this result, and compare it with the pressure predicted by the ideal gas law.

52. You want to store 165 g of CO₂ gas in a 12.5-L tank at room temperature (25 °C). Calculate the pressure the gas would have using (a) the ideal gas law and (b) the van der Waals equation. (For CO₂, \( a = 3.59 \text{ atm} \cdot \text{L}^2/\text{mol}^2 \) and \( b = 0.0427 \text{ L} \cdot \text{mol} \).

**General Questions**

These questions are not designated as to type or location in the chapter. They may combine several concepts.

53. Complete the following table:

<table>
<thead>
<tr>
<th></th>
<th>atm</th>
<th>mm Hg</th>
<th>kPa</th>
<th>bar</th>
</tr>
</thead>
<tbody>
<tr>
<td>Standard atmosphere</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Partial pressure of N₂ in the atmosphere</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Tank of compressed H₂</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Atmospheric pressure at the top of Mount Everest</td>
<td></td>
<td></td>
<td>33.7</td>
<td></td>
</tr>
</tbody>
</table>

54. On combustion, 1.0 L of a gaseous compound of hydrogen, carbon, and nitrogen gives 2.0 L of CO₂, 3.5 L of H₂O vapor, and 0.50 L of N₂ at STP. What is the empirical formula of the compound?

55. ▲ You have a sample of helium gas at –33 °C, and you want to increase the rms speed of helium atoms by 10.0%. To what temperature should the gas be heated to accomplish this?

56. If 12.0 g of O₂ is required to inflate a balloon to a certain size at 27 °C, what mass of O₂ is required to inflate it to the same size (and pressure) at 50 °C?

57. Butyl mercaptan, C₄H₉SH, has a very bad odor and is among the compounds added to natural gas to help detect a leak of otherwise odorless natural gas. In an experiment, you burn 95.0 mg of C₄H₉SH and collect the product gases (SO₂, CO₂, and H₂O) in a 5.25-L flask at 25 °C. What is the total gas pressure in the flask, and what is the partial pressure of each of the product gases?

58. A bicycle tire has an internal volume of 1.52 L and contains 0.406 mol of air. The tire will burst if its internal pressure reaches 7.25 atm. To what temperature, in degrees Celsius, does the air in the tire need to be heated to cause a blowout?

59. The temperature of the atmosphere on Mars can be as high as 27 °C at the equator at noon, and the atmospheric pressure is about 8 mm Hg. If a spacecraft could collect 10. m³ of this atmosphere, compress it to a small volume, and send it back to Earth, how many moles would the sample contain?

60. If you place 2.25 g of solid silicon in a 6.56-L flask that contains CH₃Cl with a pressure of 585 mm Hg at 25 °C, what mass of dimethylchlorosilane, \((\text{CH₃})₂\text{SiCl}_2(g)\), can be formed?

\[ \text{Si(s)} + 2 \text{CH₃Cl(g)} \rightarrow (\text{CH₃})₂\text{SiCl}_2(g) \]

What pressure of \((\text{CH₃})₂\text{SiCl}_2(g)\) would you expect in this same flask at 95 °C on completion of the reaction? (Dimethylchlorosilane is one starting material used to make silicones, polymeric substances used as lubricants, antistick agents, and water-proofing caulks.)

61. Ni(CO)₄ can be made by reacting finely divided nickel with gaseous CO. If you have CO in a 1.50-L flask at a pressure of 418 mm Hg at 25.0 °C, along with 0.450 g of Ni powder, what is the theoretical yield of Ni(CO)₄?

62. Ethane burns in air to give H₂O and CO₂.

\[ 2 \text{C}_2\text{H}_6(g) + 7 \text{O}_2(g) \rightarrow 4 \text{CO}_2(g) + 6 \text{H}_2\text{O}(g) \]

(a) Four gases are involved in this reaction. Place them in order of increasing rms speed. (Assume all are at the same temperature.)

(b) A 3.264-L flask contains C₂H₆ at a pressure of 256 mm Hg and a temperature of 25 °C. Suppose O₂ gas is added to the flask until C₂H₆ and O₂ are in the correct stoichiometric ratio for the combustion reaction. At this point, what is the partial pressure of O₂ and what is the total pressure in the flask?
63. You have four gas samples:
1. 1.0 L of H₂ at STP
2. 1.0 L of Ar at STP
3. 1.0 L of H₂ at 27 °C and 760 mm Hg
4. 1.0 L of He at 0 °C and 900 mm Hg
(a) Which sample has the largest number of gas particles (atoms or molecules)?
(b) Which sample contains the smallest number of particles?
(c) Which sample represents the largest mass?

64. Propane reacts with oxygen to give carbon dioxide and water vapor.

\[ \text{C}_3\text{H}_8(g) + 5 \text{O}_2(g) \rightarrow 3 \text{ CO}_2(g) + 4 \text{ H}_2\text{O}(g) \]

If you mix C₃H₈ and O₂ in the correct stoichiometric ratio, and if the total pressure of the mixture is 288 mm Hg, what are the partial pressures of C₃H₈ and O₂? If the temperature and volume do not change, what is the pressure of the water vapor?

65. Iron carbonyl can be made by the direct reaction of iron metal and carbon monoxide.

\[ \text{Fe}(s) + 5 \text{ CO}(g) \rightarrow \text{Fe} (\text{CO})_5(\ell) \]

What is the theoretical yield of Fe(CO)₅ if 3.52 g of iron is treated with CO gas having a pressure of 732 mm Hg in a 5.00-L flask at 23 °C?

66. Analysis of a gaseous chlorofluorocarbon, CCl₂F₂, shows that it contains 11.79% C and 69.57% Cl. In another experiment, you find that 0.107 g of the compound fills a 458-mL flask at 25 °C with a pressure of 21.3 mm Hg. What is the molecular formula of the compound?

67. There are five compounds in the family of sulfur-fluorine compounds with the general formula SF₅. One of these compounds is 25.23% S. If you place 0.0955 g of the compound in a 89-mL flask at 45 °C, the pressure of the gas is 83.8 mm Hg. What is the molecular formula of SF₅?

68. A miniature volcano can be made in the laboratory with ammonium dichromate. When ignited, it decomposes in a fiery display.

\[ (\text{NH}_4)_2\text{Cr}_2\text{O}_7(s) \rightarrow \text{N}_2(g) + 4 \text{ H}_2\text{O}(g) + \text{Cr}_2\text{O}_3(s) \]

If 0.95 g of ammonium dichromate is used and the gases from this reaction are trapped in a 15.9-L flask at 23 °C, what is the total pressure of the gas in the flask? What are the partial pressures of N₂ and H₂O?

69. The density of air 20 km above Earth's surface is 92 g/m³. The pressure of the atmosphere is 42 mm Hg, and the temperature is −63 °C.
(a) What is the average molar mass of the atmosphere at this altitude?
(b) If the atmosphere at this altitude consists of only O₂ and N₂, what is the mole fraction of each gas?

70. A 3.0-L bulb containing He at 145 mm Hg is connected by a valve to a 2.0-L bulb containing Ar at 355 mm Hg (see figure below). Calculate the partial pressure of each gas and the total pressure after the valve between the flasks is opened.

![Image of two bulbs connected by a valve](image)

71. Chlorine dioxide, ClO₂, reacts with fluorine to give a new gas that contains Cl, O, and F. In an experiment, you find that 0.150 g of this new gas has a pressure of 17.2 mm Hg in a 1850-mL flask at 21 °C. What is the identity of the unknown gas?

72. A xenon fluoride can be prepared by heating a mixture of Xe and F₂ gases to a high temperature in a pressure-proof container. Assume that xenon gas was added to a 0.25-L container until its pressure reached 0.12 atm at 0.0 °C. Fluorine gas was then added until the total pressure reached 0.72 atm at 0.0 °C. After the reaction was complete, the xenon was consumed completely, and the pressure of the F₂ remaining in the container was 0.36 atm at 0.0 °C. What is the empirical formula of the xenon fluoride?

73. Which of the following is not correct?
(a) Diffusion of gases occurs more rapidly at higher temperatures.
(b) Effusion of H₂ is faster than effusion of He (assume similar conditions and a rate expressed in units of mol/h).
(c) Diffusion will occur faster at low pressure than at high pressure.
(d) The rate of effusion of a gas (mol/h) is directly proportional to molar mass.

74. The ideal gas law is least accurate under conditions of high pressure and low temperature. In those situations, using the van der Waals equation is advisable.
(a) Calculate the pressure exerted by 12.0 g of CO₂ in a 500-mL vessel at 298 K, using the ideal gas equation. Then, recalculate the pressure using the van der Waals equation. Assuming the pressure calculated from van der Waal's equation is correct, what is the percent error in the answer when using the ideal gas equation?
(b) Next, cool this sample to −70 °C. Then perform the same calculation for the pressure exerted by CO₂ at this new temperature, using both the ideal gas law and the van der Waals equation. Again, what is the percent error when using the ideal gas equation?

75. Carbon dioxide, CO₂, was shown to effuse through a porous plate at the rate of 0.033 mol/min. The same quantity of an unknown gas, 0.033 moles, is found to effuse through the same porous barrier in 104 seconds. Calculate the molar mass of the unknown gas.

76. In an experiment, you have determined that 0.66 moles of CF₄ effuse through a porous barrier over a 4.8 minute period. How long will it take for 0.66 moles of CH₄ to effuse through the same barrier?

77. A balloon is filled with helium gas to a gauge pressure of 22 mm Hg at 25 °C. The volume of the gas is 305 mL, and the barometric pressure is 755 mm Hg. What amount of helium is in the balloon? (Remember that gauge pressure = total pressure − barometric pressure. See page 511.)

78. If you have a sample of water in a closed container, some of the water will evaporate until the pressure of the water vapor, at 25 °C, is 23.8 mm Hg. How many molecules of water per cubic centimeter exist in the vapor phase?

79. You are given 1.56 g of a mixture of KClO₃ and KCl. When heated, the KClO₃ decomposes to KCl and O₂:

\[ 2 \text{KClO}_3(s) \rightarrow 2 \text{KCl}(s) + 3 \text{O}_2(g) \]

and 327 mL of O₂ with a pressure of 735 mm Hg is collected at 19 °C. What is the weight percentage of KClO₃ in the sample?

80. A study of climbers who reached the summit of Mount Everest without supplemental oxygen showed that the partial pressures of O₂ and CO₂ in their lungs were 35 mm Hg and 7.5 mm Hg, respectively. The barometric pressure at the summit was 253 mm Hg. Assume the lung gases are saturated with moisture at a body temperature of 37 °C [which means the partial pressure of water vapor in the lungs is \( P(H_2O) = 47.1 \text{ mm Hg} \)]. If you assume the lung gases consist of only O₂, N₂, CO₂, and H₂O, what is the partial pressure of N₂?

81. Nitrogen monoxide reacts with oxygen to give nitrogen dioxide:

\[ 2 \text{NO}(g) + \text{O}_2(g) \rightarrow 2 \text{NO}_2(g) \]

(a) Place the three gases in order of increasing rms speed at 298 K.

(b) If you mix NO and O₂ in the correct stoichiometric ratio and NO has a partial pressure of 150 mm Hg, what is the partial pressure of O₂?

(c) After reaction between NO and O₂ is complete, what is the pressure of NO₂ if the NO originally had a pressure of 150 mm Hg and O₂ was added in the correct stoichiometric amount?

82. Ammonia gas is synthesized by combining hydrogen and nitrogen:

\[ 3 \text{H}_2(g) + \text{N}_2(g) \rightarrow 2 \text{NH}_3(g) \]

(a) If you want to produce 562 g of NH₃, what volume of H₂ gas, at 56 °C and 745 mm Hg, is required?

(b) Nitrogen for this reaction will be obtained from air. What volume of air, measured at 29 °C and 745 mm Hg pressure, will be required to provide the nitrogen needed to produce 562 g of NH₃? Assume the sample of air contains 78.1 mole % N₂.

83. Nitrogen trifluoride is prepared by the reaction of ammonia and fluorine:

\[ 4 \text{NH}_3(g) + 3 \text{F}_2(g) \rightarrow 3 \text{NH}_4\text{F}(s) + \text{NF}_3(g) \]

If you mix NH₃ with F₂ in the correct stoichiometric ratio, and if the total pressure of the mixture is 120 mm Hg, what are the partial pressures of NH₃ and F₂? When the reactants have been completely consumed, what is the total pressure in the flask? (Assume T is constant.)

84. Chlorine trifluoride, ClF₃, is a valuable reagent because it can be used to convert metal oxides to metal fluorides:

\[ 6 \text{NiO}(s) + 4 \text{ClF}_3(g) \rightarrow 6 \text{NiF}_2(s) + 2 \text{Cl}_2(g) + 3 \text{O}_2(g) \]

(a) What mass of NiO will react with ClF₃ gas if the gas has a pressure of 250 mm Hg at 20 °C in a 2.5-L flask?

(b) If the ClF₃ described in part (a) is completely consumed, what are the partial pressures of Cl₂ and O₂ in the 2.5-L flask at 20 °C (in mm Hg)? What is the total pressure in the flask?

85. Relative humidity is the ratio of the partial pressure of water in air at a given temperature to the vapor pressure of water at that temperature. Calculate the mass of water per liter of air under the following conditions:

(a) at 20 °C and 45% relative humidity

(b) at 0 °C and 95% relative humidity

Under which circumstances is the mass of H₂O per liter greater? (See Appendix G for the vapor pressure of water.)

86. How much water vapor is present in a dormitory room when the relative humidity is 55% and the temperature is 25 °C? The dimensions of the room are 4.5 m² floor area and 3.5 m ceiling height. (See Study Question 85 for a definition of relative humidity and Appendix G for the vapor pressure of water.)

In the Laboratory

87. You have a 550-mL tank of gas with a pressure of 1.56 atm at 24 °C. You thought the gas was pure carbon monoxide gas, CO, but you later found it was contaminated by small quantities of gaseous CO₂ and O₂. Analysis shows that the tank pressure is 1.34 atm (at 24 °C) if the CO₂ is removed. Another experiment shows that 0.0870 g of O₂ can be removed chemically. What are the masses of CO and CO₂ in the tank, and what is the partial pressure of each of the three gases at 25 °C?
88. ▲ Methane is burned in a laboratory Bunsen burner to give CO₂ and water vapor. Methane gas is supplied to the burner at the rate of 5.0 L/min (at a temperature of 28 °C and a pressure of 775 mm Hg). At what rate must oxygen be supplied to the burner (at a pressure of 742 mm Hg and a temperature of 26 °C)?

89. ▲ Iron forms a series of compounds of the type Fe₅(CO)₄. In air, these compounds are oxidized to Fe₂O₃ and CO₂ gas. After heating a 0.142-g sample of Fe₅(CO)₄, in air, you isolate the CO₂ in a 1.50-L flask at 25 °C. The pressure of the gas is 44.9 mm Hg. What is the empirical formula of Fe₅(CO)₄?

90. ▲ Group 2A metal carbonates are decomposed to the metal oxide and CO₂ on heating:

MCO₃(s) → MO(s) + CO₂(g)

You heat 0.158 g of a white, solid carbonate of a Group 2A metal (M) and find that the evolved CO₂ has a pressure of 69.8 mm Hg in a 285-mL flask at 25 °C. Identify M.

91. One way to synthesize diborane, B₂H₆, is the reaction

2 NaBH₄(s) + 2 H₃PO₄(ℓ) → B₂H₆(g) + 2 NaH₂PO₄(s) + 2 H₂(g)

(a) If you have 0.136 g of NaBH₄ and excess H₃PO₄,
and you collect the B₂H₆ in a 2.75-L flask at 25 °C,
what is the pressure of the B₂H₆ in the flask?

(b) A by-product of the reaction is H₂ gas. If both B₂H₆
and H₂ gas come from this reaction, what is the
total pressure in the 2.75-L flask (after reaction of
0.136 g of NaBH₄ with excess H₃PO₄) at 25 °C?

92. You are given a solid mixture of NaNO₂ and NaCl and are asked to analyze it for the amount of NaNO₂ present. To do so, you allow the mixture to react with sulfamic acid, HSO₃NH₂, in water according to the equation

NaNO₂(aq) + HSO₃NH₂(aq) → NaHSO₄(aq) + H₂O(ℓ) + N₂(g)

What is the weight percentage of NaNO₂ in 1.232 g of the solid mixture if reaction with sulfamic acid produces 295 mL of dry N₂ gas with a pressure of 713 mm Hg at 21.0 °C?

93. ▲ You have 1.249 g of a mixture of NaHCO₃ and Na₂CO₃. You find that 12.0 mL of 1.50 M HCl is required to convert the sample completely to NaCl, H₂O, and CO₂.

NaHCO₃(aq) + HCl(aq) → NaCl(aq) + H₂O(ℓ) + CO₂(g)

Na₂CO₃(aq) + 2 HCl(aq) → 2 NaCl(aq) + H₂O(ℓ) + CO₂(g)

What volume of CO₂ is evolved at 745 mm Hg and 25 °C?

94. ▲ A mixture of NaHCO₃ and Na₂CO₃ has a mass of 2.50 g. When treated with HCl(aq), 665 mL of CO₂ gas is liberated with a pressure of 735 mm Hg at 25 °C. What is the weight percent of NaHCO₃ and Na₂CO₃ in the mixture? (See Study Question 93 for the reactions that occur.)

95. ▲ Many nitrate salts can be decomposed by heating. For example, blue, anhydrous copper(II) nitrate produces the gases nitrogen dioxide and oxygen when heated. In the laboratory, you find that a sample of this salt produced a 0.195-g mixture of gaseous NO₂ and O₂ with a total pressure of 725 mm Hg at 35 °C in a 125-mL flask (and black, solid CuO was left as a residue). What is the average molar mass of the gas mixture? What are the mole fractions of NO₂ and O₂ in the mixture? What amount of each gas is in the mixture? Do these amounts reflect the relative amounts of NO₂ and O₂ expected based on the balanced equation? Is it possible that the fact that some NO₂ molecules combine to give N₂O₄ plays a role?

96. ▲ A compound containing C, H, N, and O is burned in excess oxygen. The gases produced by burning 0.1152 g are first treated to convert the nitrogen-containing product gases into N₂, and then the resulting mixture of CO₂, H₂O, N₂, and excess O₂ is passed through a bed of CaCl₂ to absorb the water. The CaCl₂ increases in mass by 0.09912 g. The remaining gases are bubbled into water to form H₂CO₃, and this solution is titrated with 0.3283 M NaOH; 28.81 mL is required to achieve the second equivalence point. The excess O₂ gas is removed by reaction with copper metal (to give CuO). Finally, the N₂ gas is collected in a 225.0-mL flask, where it has a pressure of 65.12 mm Hg at 25 °C. In a separate experiment, the unknown compound is found to have a molar mass of 150 g/mol. What are the empirical and molecular formulas of the unknown compound?
97. You have a gas, one of the three known phosphorus–fluorine compounds (PF₃, PF₅, and P₂F₆). To find out which, you have decided to measure its molar mass.  
(a) First, you determine that the density of the gas is 5.60 g/L at a pressure of 0.971 atm and a temperature of 18.2 °C. Calculate the molar mass and identify the compound.  
(b) To check the results from part (a), you decide to measure the molar mass based on the relative rates of effusion of the unknown gas and CO₂. You find that CO₂ effuses at a rate of 0.050 mol/min, whereas the unknown phosphorus fluoride effuses at a rate of 0.028 mol/min. Calculate the molar mass of the unknown gas based on these results.

98. A 1.50 L constant volume calorimeter (Figure 5.12) contains C₃H₆(g) and O₂(g). The partial pressure of C₃H₆ is 0.10 atm and the partial pressure of O₂ is 5.0 atm. The temperature is 20.0 °C. A reaction occurs between the two compounds, forming CO₂(g) and H₂O(ℓ). The heat from the reaction causes the temperature to rise to 23.2 °C.  
(a) Write a balanced chemical equation for the reaction.  
(b) How many moles of C₃H₆(g) are present in the flask initially?  
(c) What is the mole fraction of C₃H₆(g) in the flask before reaction?  
(d) After the reaction, the flask contains excess oxygen and the products of the reaction, CO₂(g) and H₂O(ℓ). What amount of unreacted O₂(g) remains?  
(e) After the reaction, what is the partial pressure exerted by the CO₂(g) in this system?  
(f) What is the partial pressure exerted by the excess oxygen remaining after the reaction?

Summary and Conceptual Questions
The following questions may use concepts from this and previous chapters.

99. A 1.0-L flask contains 10.0 g each of O₂ and CO₂ at 25 °C.  
(a) Which gas has the greater partial pressure, O₂ or CO₂, or are they the same?  
(b) Which molecules have the greater rms speed, or are they the same?  
(c) Which molecules have the greater average kinetic energy, or are they the same?

100. If equal masses of O₂ and N₂ are placed in separate containers of equal volume at the same temperature, which of the following statements is true? If false, explain why it is false.  
(a) The pressure in the flask containing N₂ is greater than that in the flask containing O₂.  
(b) There are more molecules in the flask containing O₂ than in the flask containing N₂.

101. You have two pressure-proof steel cylinders of equal volume, one containing 1.0 kg of CO and the other containing 1.0 kg of acetylene, C₂H₂.  
(a) In which cylinder is the pressure greater at 25 °C?  
(b) Which cylinder contains the greater number of molecules?

102. Two flasks, each with a volume of 1.00 L, contain O₂ gas with a pressure of 380 mm Hg. Flask A is at 25 °C, and flask B is at 0 °C. Which flask contains the greater number of O₂ molecules?  
(a) You have 65.0 g of sodium, a 35.0-L flask containing Na₂O gas with a pressure of 2.12 atm at 23 °C, and excess ammonia. What is the theoretical yield (in grams) of NaN₃?  
(b) Draw a Lewis structure for the azide ion. Include all possible resonance structures. Which resonance structure is most likely?  
(c) What is the shape of the azide ion?
107. If the absolute temperature of a gas doubles, by how much does the rms speed of the gaseous molecules increase?

108. ▲ Chlorine gas \((\text{Cl}_2)\) is used as a disinfectant in municipal water supplies, although chlorine dioxide \((\text{ClO}_2)\) and ozone are becoming more widely used. \text{ClO}_2 is a better choice than \text{Cl}_2 in this application because it leads to fewer chlorinated by-products, which are themselves pollutants.
(b) How many valence electrons are in \text{ClO}_2? (c) What is the hybridization of the central Cl atom in \text{ClO}_2\(^-\)? What is the shape of the ion?
(d) Which species has the larger bond angle, \text{O}_3 or \text{ClO}_2\(^-\)? Explain briefly.
(e) Chlorine dioxide, \text{ClO}_2, a yellow-green gas, can be made by the reaction of chlorine with sodium chlorite:

\[
2 \text{NaClO}_2(s) + \text{Cl}_2(g) \rightarrow 2 \text{NaCl(s)} + 2 \text{ClO}_2(g)
\]
Assume you react 15.6 g of \text{NaClO}_2 with chlorine gas, which has a pressure of 1050 mm Hg in a 1.45-L flask at 22 °C. What mass of \text{ClO}_2 can be produced?
Applying Chemical Principles

The Goodyear Blimp

Goodyear blimps are a familiar sight at sporting events. They hover at altitudes between 1000–3000 feet, providing a stable platform for television cameras. The outer surface, or envelope, of a Goodyear blimp is constructed of polyester fabric impregnated with neoprene rubber. The envelope contains the lighter-than-air gas helium. Within the fore and aft sections of the envelope are two air-filled structures called ballonets, which serve two purposes. As the blimp changes altitude, the ballonets are inflated or deflated with air to maintain a helium pressure within the envelope that is similar to the external air pressure. The ballonets also are used to keep the craft level. The pilot and passengers ride in the gondola, which is attached to the bottom of the envelope.

One of the Goodyear blimps has a gross weight of 12,840 lbs (5820 kg), and the volume of the inner envelope is 202,700 ft³ (5740 m³). The ballonets, when both fully inflated with air, occupy a total of 18,000 ft³ (510 m³) of the envelope. Helium is neither added nor removed from the envelope during operation. Loss of helium through the rubber impregnated polyester fabric is slow, at approximately 10,000 ft³ (280 m³) per month.

Questions:

1. At sea level, atmospheric pressure is 1.00 atm. Calculate the density (in g/L) of helium at this pressure and 25 °C.

2. Use the data provided in Table 11.1 to calculate the density of dry air at 1.00 atm and 25 °C.

3. To stay aloft, a blimp must achieve neutral buoyancy; that is, its density must equal that of the surrounding air. The density of the blimp is its total weight (blimp, helium and air, passengers, and ballast) divided by its volume. Assume that the gross weight of the blimp includes the blimp’s structure and the helium, but does not include the air in the ballonets or the weights of the passengers and ballast. If the ballonets are filled with 12,000 ft³ (340 m³) of air at 1.00 atm and 25 °C, what additional weight (of passengers and ballast) is required for neutral buoyancy?