States of Matter

Section 13.1 Gases

pages 385–392

Practice Problems

pages 388, 392

1. Calculate the ratio of effusion rates for nitrogen (N_2) and neon (Ne).

$$\frac{\text{Rate}_{\text{nitrogen}}}{\text{Rate}_{\text{neon}}} = \sqrt{\frac{20.2\ \cancel{g/mol}}{28.0\ \cancel{g/mol}}} = \sqrt{0.721} = 0.849$$

2. Calculate the ratio of diffusion rates for carbon monoxide (CO) and carbon dioxide (CO_2).

$$\frac{\text{Rate}_{\text{carbon monoxide}}}{\text{Rate}_{\text{carbon dioxide}}} = \sqrt{\frac{44.0\ \cancel{g/mol}}{28.0\ \cancel{g/mol}}} = \sqrt{1.57}$$

$= 1.25$

3. What is the rate of effusion for a gas that has a molar mass twice that of a gas that effuses at a rate of 3.6 mol/min?

Rearrange Graham's law to solve for Rate_A.

$$\text{Rate}_A = \text{Rate}_B \times \sqrt{\frac{\text{molar mass}_B}{\text{molar mass}_A}}$$

$\text{Rate}_B = 3.6\ \text{mol/min}$

$$\frac{\text{molar mass}_B}{\text{molar mass}_A} = 0.5$$

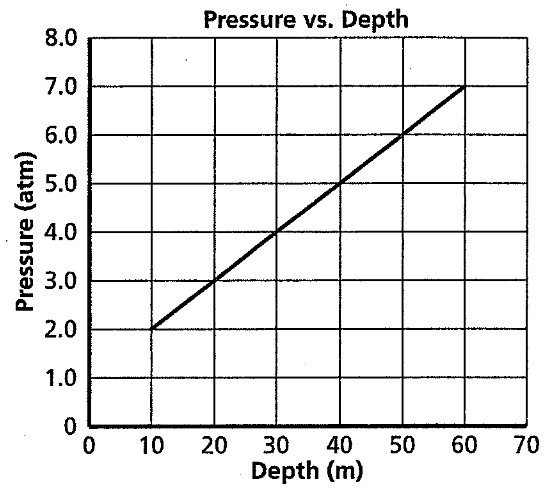
$\text{Rate}_A = 3.6\ \text{mol/min} \times \sqrt{0.5}$

$= 3.6\ \text{mol/min} \times 0.71$

$= 2.5\ \text{mol/min}$

Problem-Solving Lab

page 390



4. What is the partial pressure of hydrogen gas in a mixture of hydrogen and helium if the total pressure is 600 mm Hg and the partial pressure of helium is 439 mm Hg?

$P_{\text{hydrogen}} = P_{\text{total}} - P_{\text{helium}}$

$= 600\ \text{mm Hg} - 439\ \text{mm Hg}$

$= 161\ \text{mm Hg}$

5. Find the total pressure for a mixture that contains four gases with partial pressures of 5.00 kPa, 4.56 kPa, 3.02 kPa, and 1.20 kPa.

$P_{\text{total}} = 5.00\ \text{kPa} + 4.56\ \text{kPa} + 3.02\ \text{kPa} + 1.20\ \text{kPa}$

$= 13.78\ \text{kPa}$

6. Find the partial pressure of carbon dioxide in a gas mixture with a total pressure of 30.4 kPa if the partial pressures of the other two gases in the mixture are 16.5 kPa and 3.7 kPa.

$P_{\text{carbon dioxide}} = 30.4\ \text{kPa} - (16.5\ \text{kPa} + 3.7\ \text{kPa})$

$= 30.4\ \text{kPa} - 20.2\ \text{kPa} = 10.2\ \text{kPa}$

Section 13.1 Assessment

page 392

7. What assumption of the kinetic-molecular theory explains why a gas can expand to fill a container?

the assumption that gas particles are in constant and random motion

8. How does the mass of a gas particle affect its rate of effusion?

The rate of effusion decreases as mass increases.

9. Suppose two gases in a container have a total pressure of 1.20 atm. What is the pressure of gas B if the partial pressure of gas A is 0.75 atm?

$P_B = P_{total} - P_A = 1.20 \text{ atm} - 0.75 \text{ atm} = 0.45 \text{ atm}$

10. Explain how changes in atmospheric pressure affect the height of the column of mercury in a barometer.

An increase in atmospheric pressure increases the pressure on the surface of mercury and the column rises; a decrease in atmospheric pressure decreases the pressure on the surface of the mercury and the column drops.

11. Recognizing Cause and Effect Explain why a tire or balloon expands when air is added.

When air is added, there are more collisions between air particles and the walls, exerting greater pressure inside the balloon or tire and expanding the walls.

12. Thinking Critically Explain why the container of water must be inverted when a gas is collected by displacement of water.

If the container is not inverted, the gas, which is less dense than water, will rise through the water and escape from the opening of the container.

Section 13.2 Forces of Attraction

pages 393–395

Section 13.2 Assessment

page 395

13. Why are dipole–dipole forces stronger than dispersion forces for molecules of comparable mass?

Dipole-dipole forces exist between permanent dipoles; dispersion forces exist between temporary dipoles.

14. Which of the molecules listed below can form hydrogen bonds? For which of the molecules would dispersion forces be the only intermolecular force? Give reasons for your answers.

a. H_2

b. NH_3

c. HCl

d. HF

hydrogen bonds: b, d; only dispersion forces: a; b and d are polar molecules with a highly electronegative atom bonded to hydrogen. a is nonpolar.

15. Predicting Make a prediction about the relative boiling points of the noble gases. Give a reason for your answer.

The boiling points will increase from helium to xenon because larger noble gas atoms have more electrons, making the dispersion forces between those atoms stronger. It takes more energy to overcome those forces.

16. Thinking Critically In a methane molecule (CH_4), there are 4 single covalent bonds. In an octane molecule (C_8H_{18}), there are 25 single covalent bonds. How does the number of bonds affect the dispersion forces in samples of methane and octane? Which compound is a gas at room temperature? Which is a liquid?

More bonds mean more electrons to form temporary dipoles, which means greater dispersion forces. Methane is a gas; octane is a liquid.

Section 13.3 Liquids and Solids

pages 396–403

Section 13.3 Assessment

page 403

17. Explain how hydrogen bonds affect the viscosity of a liquid. How does a change in temperature affect viscosity?

Hydrogen bonds increase viscosity because they increase the liquid's intermolecular forces, making it more resistant to flow. Viscosity decreases with temperature.

18. What effect does soap have on the surface tension of water?

Soaps and detergents decrease the surface tension of water.

19. How are a unit cell and a crystal lattice related?

Unit cells are the building blocks of the crystal lattice.

20. Explain why solids are not classified as fluids.

The particles in a solid only vibrate around fixed locations because they are held in place by strong attractive forces.

21. What is the difference between a molecular solid and a covalent network solid?

A molecular solid is formed from molecules held together by intermolecular forces that are weaker than covalent bonds; covalent network solids are formed from molecules held together by covalent bonds.

22. Explain why most solids are denser than most liquids at the same temperature.

Solids are usually denser than liquids because their particles are held more closely together by intermolecular attractions.

23. Thinking Critically Hypothesize why the surface of mercury in a thermometer is convex; that is, the surface is higher at the center.

The cohesion between mercury atoms is stronger than the adhesion between mercury and glass.

Section 13.4 Phase Changes

pages 404–409

Section 13.4 Assessment

page 409

24. What information does a phase diagram supply?

the combinations of temperature and pressure under which a given substance exists as a solid, liquid, and gas

25. What is the major difference between the processes of melting and freezing?

Freezing occurs when a liquid becomes a solid and energy is released. Melting requires an input of energy to convert a solid to a liquid.

26. Explain what the triple point and the critical point on a phase diagram represent.

triple point: the temperature at which the three phases of a substance can coexist; critical point: the pressure and temperature above which a substance cannot exist as a liquid

27. Comparing and Contrasting Compare what happens to the energy, order, and spacing of particles when a solid other than ice changes to a liquid with what happens to the energy, order, and spacing of particles when a gas changes to a liquid.

When a solid changes to a liquid, the particles gain kinetic energy, become less ordered, and are spaced farther apart. When a gas changes to a liquid, the particles lose kinetic energy, become more ordered, and are spaced closer together.

28. Thinking Critically Aerosol cans contain compressed gases that, when released, help propel the contents out of the can. Why is it important to keep aerosol cans from overheating?

The gas inside an aerosol can is compressed into a small space under pressure. Upon heating, the kinetic energy of the gas particles increases, raising the pressure. The can will explode.

CHEMLAB
page 411

Analyze and Conclude

3. **Interpreting Data** Make a generalization about the shape of a liquid drop and the evaporation rate of the liquid.

   The more spread out the liquid drop, the higher the evaporation rate; the more spherical or dome-shaped the drop, the lower the rate.

4. **Recognizing Cause and Effect** What is the relationship between surface tension and the shape of a liquid drop? What are the attractive forces that increase surface tension?

   The greater the surface tension, the more spherical or dome-shaped the drop. Intermolecular forces increase surface tension.

5. **Applying Concepts** The isopropyl alcohol you used is a mixture of isopropyl alcohol and water. Would pure isopropyl alcohol evaporate more quickly or more slowly compared to the alcohol and water mixture? Give a reason for your answer.

   Because of less hydrogen bonding, pure alcohol would evaporate more quickly than a mixture of alcohol and water.

6. **Thinking Critically** Household ammonia is a mixture of ammonia and water. Based on the data you collected, is there more ammonia or more water in the mixture? Use what you learned about the relative strengths of the attractive forces in ammonia and water to support your conclusion.

   More water; the evaporation rate is more similar to that of water. (The hydrogen bonds in ammonia are weaker than those in water; thus, ammonia is a gas at room temperature.)

7. **Drawing a Conclusion** How does the rate of evaporation of warm ethanol compare to ethanol at room temperature? Use kinetic-molecular theory to explain your observations.

   Warm ethanol evaporates faster than ethanol at room temperature. The average kinetic energy of the ethanol molecules increases at the higher temperature. Thus, a larger number of ethanol molecules have the energy needed to escape the surface of the liquid and evaporate.

Real-World Chemistry

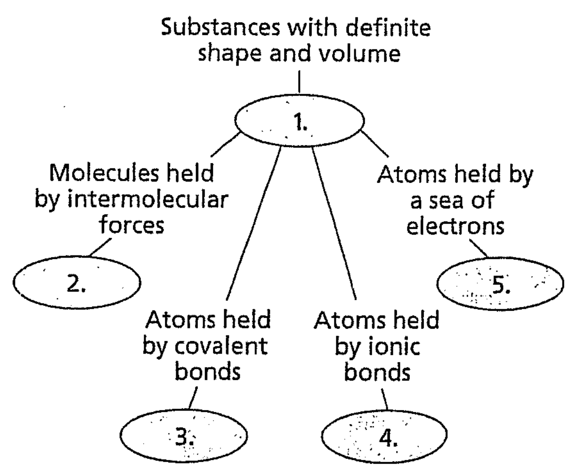
2. Suggest why a person who has a higher than normal temperature might be given a rubdown with rubbing alcohol (70% isopropyl alcohol).

   Rubbing alcohol has a high evaporation rate. Since evaporation requires energy, applying it to the skin can cause absorption of heat energy and a lowering of body temperature.

Chapter 13 Assessment *pages 414–416*

Concept Mapping

29. Complete the concept map using the following terms: covalent network solid, molecular solid, metallic solid, ionic solid, solid.



1. solid; 2. molecular solid; 3. covalent network solid; 4. ionic solid; 5. metallic solid

Mastering Concepts

30. What is an elastic collision? (13.1)

one in which no kinetic energy is lost

31. How does the kinetic energy of particles vary as a function of temperature? (13.1)

It is directly proportional to the temperature.

32. Use the kinetic-molecular theory to explain the compression and expansion of gases. (13.1)

Because of the space between gas particles, gases are easily compressed when pushed into a smaller volume. When the pressure is removed, their random motion enables gases to expand.

33. Compare diffusion and effusion. Explain the relationship between the rates of these processes and the molar mass of a gas. (13.1)

Both involve the movement of particles. Diffusion is the movement of one substance through another; effusion occurs when a substance under pressure escapes through a tiny opening. Effusion and diffusion rates are inversely related to molecular mass of a gas.

34. What happens to the density of gas particles in a cylinder as the piston is raised? (13.1)

Density decreases because the gas particles occupy more volume per unit mass.

35. Explain why the baking instructions on a box of cake mix are different for high and low elevations. Would you expect to have a longer or shorter cooking time at a high elevation? (13.1)

because of the variation in air pressure with elevation; At high elevations, reduced air pressure results in a lower boiling point for water and cooking time is longer.

36. Explain the difference between a temporary dipole and a permanent dipole. (13.2)

A temporary dipole forms when electrons move toward the near side of one nonpolar particle and the far side of another to avoid repelling each other. A permanent dipole is found only in polar molecules in which the difference in electronegativity between two atoms causes the electron orbits to shift toward the most electronegative atom.

37. Why are dispersion forces weaker than dipole–dipole forces? (13.2)

Dispersion forces exist between temporary dipoles; dipole-dipole forces exist between permanent dipoles.

38. Explain why hydrogen bonds are stronger than most dipole–dipole forces. (13.2)

A hydrogen bond involves a large difference in electronegativity between the hydrogen atom and the atom it is attached to (O, N, or F), making the bond extremely polar.

39. Use relative differences in electronegativity to label the ends of the polar molecules listed as partially positive or partially negative. (13.2)

a. HF

$\delta+$ $\delta-$
H—F

b. NO

$\delta+$ $\delta-$
N—O

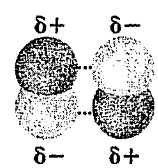
c. HBr

$\delta+$ $\delta-$
H—Br

d. CO

$\delta+$ $\delta-$
C—O

40. Draw the structure of the dipole–dipole interaction between two molecules of carbon monoxide. (13.2)



41. Decide which of the substances listed can form hydrogen bonds. (13.2)

a. H_2O
b. HF
c. NaF
d. NO
e. H_2O_2
f. NH_3
g. H_2
h. CH_4

a, b, e, f

42. Hypothesize why long, nonpolar molecules would interact more strongly with one another than spherical nonpolar molecules of similar composition. (13.2)

Because long molecules have greater surface areas, more intermolecular forces can exist.

43. What is surface tension and what conditions must exist for it to occur? (13.3)

The energy needed to increase the surface area of a liquid by a given amount; between the liquid particles, strong intermolecular forces exist.

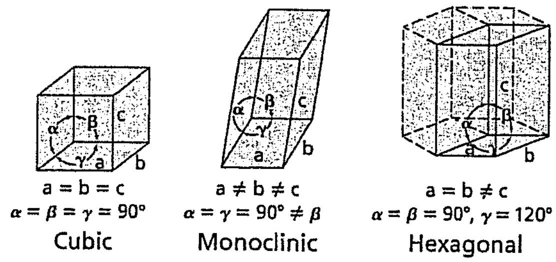
44. Explain why the surface of water in a graduated cylinder is curved. (13.3)

Adhesion between water and glass is greater than cohesion between water molecules.

45. Which liquid is more viscous at room temperature, water or molasses? Explain. (13.3)

Molasses; stronger intermolecular forces the molasses from flowing easily.

46. Use these drawings to compare the cubic, monoclinic, and hexagonal crystal systems. (13.3)



Cubic: equal sides and angles
Monoclinic: unequal sides; two equal angles
Hexagonal: two equal sides; two angles = 90°; one angle = 120°

47. What is the difference between a network solid and an ionic solid? (13.3)

Network solid is held together by covalent bonds; ionic solid by ionic bonds.

48. Explain why most metals bend when struck but most ionic solids shatter. (13.3)

The sea of electrons that holds metal ions together can easily move to accommodate outside forces; in ionic solids a strong force can separate the solid along the plane bonding groups of atoms together.

49. What is an amorphous solid? Under what conditions is such a solid likely to form? (13.3)

One that lacks a regularly repeating structure; it can form when a liquid solidifies too quickly for crystal formation to occur.

50. List the types of crystalline solids that are usually good conductors of heat and electricity. (13.3)

metallic solids; ionic solids when molten or dissolved in an aqueous solution

51. How does the strength of a liquid's intermolecular forces affect its viscosity? (13.3)

Stronger intermolecular forces result in higher viscosity because the forces hold the particles together too tightly for them to flow easily.

52. Explain why water has a higher surface tension than benzene, whose molecules are nonpolar. (13.3)

Surface tension increases with strength of interparticle forces. Water molecules are held together by strong hydrogen bonds, resulting in higher surface tension; the weaker dispersion forces between benzene molecules result in lower surface tension.

53. How does sublimation differs from deposition? (13.4)

Sublimation occurs when a solid is converted to a gas; deposition occurs when a gas is converted to a solid.

54. Compare boiling and evaporation. (13.4)

Evaporation: conversion of a liquid to a gas at the liquid's surface; boiling: when vapor pressure is equal to external atmospheric pressure; it occurs at and below the surface where bubbles form.

55. Define melting point. (13.4)

the temperature at which the crystal lattice of a solid disintegrates and it becomes a liquid

56. Explain the relationships among vapor pressure, atmospheric pressure, and boiling point. (13.4)

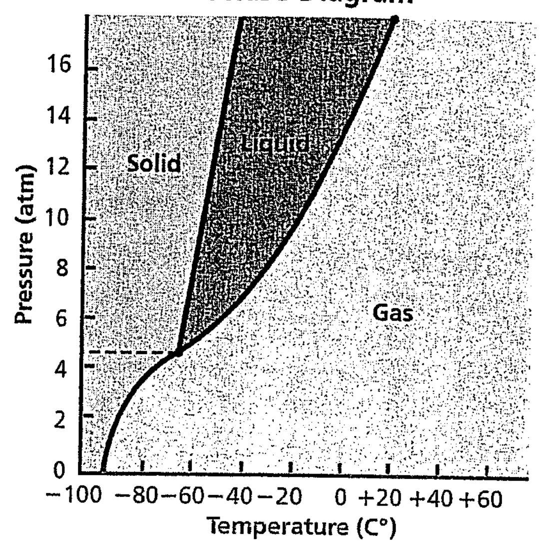
Boiling point is the temperature at which vapor pressure exerted by liquid molecules escaping from the surface of a sample equals the atmospheric pressure on the surface of the liquid.

57. Explain why dew forms on cool mornings. (13.4)

When water vapor in the air comes in contact with a cool object such as a windshield, it condenses on the object.

58. Label the solid, liquid, and gas phases, triple point, and critical point on the phase diagram shown. (13.4)

Phase Diagram



59. Why does it take more energy to boil 10 g of liquid water than to melt an equivalent mass of ice? (13.4)

Melting does not require as much energy because the particles in a solid do not have to move far apart or gain much movement to form a liquid. Gas particles are much further apart and move much more rapidly than liquid particles.

60. Why does a pile of snow slowly shrink even on days when the temperature never rises above the freezing point of water? (13.4)

Some of the snow sublimes.

61. Examine the phase diagram for carbon dioxide in **Figure 13-28.** Notice that at 1 atm pressure, the solid sublimes to a gas. What happens to the solid at much higher pressures? (13.4)

At pressures much higher than 1 atm, solid CO_2 melts.

Mastering Problems

Graham's Law of Effusion (13.1)

Level 1

62. What is the molar mass of a gas that takes three times longer to effuse than helium?

$$\frac{\text{Rate}_{gas}}{\text{Rate}_{helium}} = \sqrt{\frac{\text{mass}_{helium}}{\text{mass}_{gas}}}$$

$(\text{Rate}_{gas}/\text{Rate}_{helium})^2 = (m_{helium}m_{gas})$

$m_{gas} = (\text{Rate}_{gas}/\text{Rate}_{helium})^2(m_{helium})$

$= (3)^2(4.00\ \text{g/mol}) = 36.0\ \text{g/mol}$

Level 2

63. What is the ratio of effusion rates of krypton and neon at the same temperature and pressure?

$$\frac{\text{Rate}_{krypton}}{\text{Rate}_{neon}} = \sqrt{\frac{\text{mass}_{neon}}{\text{mass}_{krypton}}} = \sqrt{\frac{20.18\ \text{g/mol}}{83.80\ \text{g/mol}}}$$

$= 0.4931$

64. Calculate the molar mass of a gas that diffuses three times faster than oxygen under similar conditions.

$$\frac{\text{Rate}_{gas}}{\text{Rate}_{oxygen}} = \sqrt{\frac{\text{mass}_{oxygen}}{\text{mass}_{gas}}}$$

$(\text{Rate}_{gas}/\text{Rate}_{oxygen})^2 = m_{oxygen}/m_{gas}$

$m_{gas} = (\text{Rate}_{gas}/\text{Rate}_{oxygen})^2(m_{oxygen})$

$= (1/3)^2(32.00\ \text{g/mol}) = 3.56\ \text{g/mol}$

Dalton's Law of Partial Pressures (13.1)

Level 1

65. What is the partial pressure of water vapor in an air sample when the total pressure is 1.00 atm, the partial pressure of nitrogen is 0.79 atm, the partial pressure of oxygen is 0.20 atm, and the partial pressure of all other gases in air is 0.0044 atm?

$P_{total} = P_{water} + P_{all\ other\ gases}$

$P_{water} = P_{total} - P_{all\ other\ gases}$

$= 1.00\ \text{atm} - (0.20\ \text{atm} + 0.79\ \text{atm} + 0.0044\ \text{atm})$

$= 0.01\ \text{atm}$

66. What is the total gas pressure in a sealed flask that contains oxygen at a partial pressure of 0.41 atm and water vapor at a partial pressure of 0.58 atm?

$P_{total} = P_{oxygen} + P_{water\ vapor}$

$= 0.41\ \text{atm} + 0.58\ \text{atm} = 0.99\ \text{atm}$

67. Find the partial pressure of oxygen in a sealed vessel that has a total pressure of 2.6 atm and also contains carbon dioxide at 1.3 atm and helium at 0.22 atm.

$P_{total} = P_{oxygen} + P_{carbon\ dioxide} + P_{helium}$

$P_{oxygen} = P_{total} - (P_{carbon\ dioxide} + P_{helium})$

$= 2.6\ \text{atm} - (1.3\ \text{atm} + 0.22\ \text{atm}) = 1.1\ \text{atm}$

Converting Pressure Units (13.1)

Level 1

68. What is the total pressure in atmospheres of a mixture of three gases with partial pressures of 12.0 kPa, 35.6 kPa, and 22.2 kPa?

$P_{total} = 12.0\ \text{kPa} + 35.6\ \text{kPa} + 22.2\ \text{kPa} = 69.86\ \text{kPa}$

69. The pressure atop the world's highest mountain, Mount Everest, is usually about 33.6 kPa. Convert the pressure to atmospheres. How does the pressure compare with the pressure at sea level?

$$86\ \text{kPa} \times \frac{1\ \text{atm}}{101.3\ \text{kPa}} = 0.689\ \text{atm}$$

$$33.6\ \text{kPa} \times \frac{1\ \text{atm}}{101.3\ \text{kPa}} = 0.332\ \text{atm}$$

about one-third sea-level atmospheric pressure

Level 2

70. The atmospheric pressure in Denver, Colorado, is usually about 84.0 kPa. What is this pressure in atm and torr units?

$$84.0\ \text{kPa} \times \frac{1\ \text{atm}}{101.3\ \text{kPa}} = 0.829\ \text{atm}$$

$$0.829\ \text{atm} \times \frac{760\ \text{torr}}{1\ \text{atm}} = 6.30 \times 10^2\ \text{torr}$$

71. At an ocean depth of 250 feet, the pressure is about 8.4 atm. Convert the pressure to mm Hg and kPa units.

$$8.4\ \cancel{\text{atm}} \times \frac{760\ \text{mm Hg}}{1\ \cancel{\text{atm}}} = 6400\ \text{mm Hg}$$

$$8.4\ \cancel{\text{atm}} \times 101.3\ \text{kPa}/1\ \cancel{\text{atm}} = 850\ \text{kPa}$$

Mixed Review

Sharpen your problem solving skills by answering the following.

72. Use the kinetic-molecular theory to explain why both gases and liquids are fluids.

Because their particles are held together by fewer attractive forces than in solids, the particles can flow.

73. Use intermolecular forces to explain why oxygen is a gas at room temperature and water is a liquid.

Oxygen molecules are nonpolar and held together by dispersion forces, making them easy to separate. Water molecules are held together by stronger hydrogen bonds, making them harder to separate. As a result, water has a higher boiling point.

74. Use the kinetic-molecular theory to explain why gases are easier to compress than liquids or solids.

Because particles in gases are farther apart than particles in liquids or solids, there is more space for the particles to be compressed.

75. The density of mercury at 25°C and a pressure of 760 mm Hg is 13.5 g/mL; water at the same temperature and pressure has a density of 1.00 g/mL. Explain this difference in terms of intermolecular forces and the kinetic-molecular theory.

Metallic bonds holding mercury atoms together are stronger than the hydrogen bonds holding water molecules together, so the mercury atoms are more closely packed resulting in greater mass per unit volume.

76. Two flasks of equal size are connected by a narrow tube that is closed in the middle with a stopcock. One flask has no gas particles; the other flask contains 0.1 mol of hydrogen gas at a pressure of 2.0 atm.

a. Describe what happens to the gas molecules after the stopcock is opened.

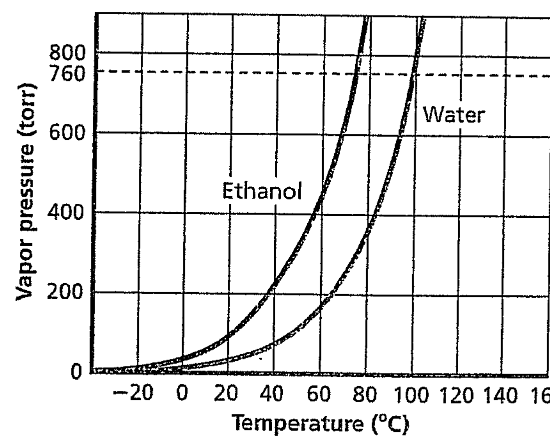
Gas will diffuse from the flask on the right into the evacuated flask until an equal number of molecules is in each flask.

b. What will happen to the gas pressure after the stopcock is opened?

It will decrease in the flask that contained H_2 and increase in the empty flask until pressure is equalized at 1.0 atm.

Thinking Critically

77. Interpreting Graphs Examine the graph below, which plots vapor pressure versus temperature for water and ethyl alcohol.



a. What is the boiling point of water at 1 atm?

100°C

b. What is the boiling point of ethyl alcohol at 1 atm?

78.5°C

c. Describe the relationship between temperature and vapor pressure for water and alcohol.

the higher the temperature, the higher the vapor pressure