

# Chapter 8 States of Matter

## Section 1 “Gases”

### Student learning outcomes-

- Describe the different states of matter and its transformation according to the kinetic molecular theory while explaining the differences between them in terms of the forces between atoms, molecules, and ions.
- Use the kinetic molecular theory to explain the properties and behaviour of gases.
- Employ gas laws to describe the behaviour of gases and their practical applications
- Develop a method (practical experiment and simulation software) to describe the different states of matter.

### New vocabulary-

- **kinetic energy:** energy due to motion
- **kinetic-molecular theory** describes the behavior of matter in terms of particles in motion.
- **Elastic collision:** is one in which no kinetic energy is lost. Kinetic energy can be transferred between colliding particles, but the total kinetic energy of the two particles does not change.
- **Temperature:** is a measure of the average kinetic energy of the particles in a sample of matter.
- **Diffusion:** is the term used to describe the movement of one material through another.
- **Graham’s law of effusion:** states that the rate of effusion for a gas is inversely proportional to the square root of its molar mass.

#### Graham’s Law

$$\text{Rate of effusion} \propto \frac{1}{\sqrt{\text{molar mass}}}$$

The rate of diffusion or effusion of a gas is inversely proportional to the square root of its molar mass.

- **Pressure:** is defined as force per unit area.
- **Barometer:** is an instrument used to measure atmospheric pressure.
- **Pascal:** is equal to a force of one newton per square meter.

- **Atmosphere:** is equal to 760 mmHg or 760 torr or 101.3 kilopascals (kPa).
- **Dalton's law of partial pressures:** states that the total pressure of a mixture of gases is equal to the sum of the pressures of all the gases in the mixture.

## Worked examples –

### EXAMPLE 1

**GRAHAM'S LAW** Ammonia has a molar mass of 17.0 g/mol; hydrogen chloride has a molar mass of 36.5 g/mol. What is the ratio of their diffusion rates?

#### 1 ANALYZE THE PROBLEM

You are given the molar masses for ammonia and hydrogen chloride. To find the ratio of the diffusion rates for ammonia and hydrogen chloride, use the equation for Graham's law of effusion.

##### Known

molar mass<sub>HCl</sub> = 36.5 g/mol

molar mass<sub>NH<sub>3</sub></sub> = 17.0 g/mol

##### Unknown

ratio of diffusion rates = ?

#### 2 SOLVE FOR THE UNKNOWN

$$\frac{\text{Rate}_{\text{NH}_3}}{\text{Rate}_{\text{HCl}}} = \sqrt{\frac{\text{molar mass}_{\text{HCl}}}{\text{molar mass}_{\text{NH}_3}}}$$

State the ratio derived from Graham's law.

$$= \sqrt{\frac{36.5 \text{ g/mol}}{17.0 \text{ g/mol}}} = 1.47$$

Substitute molar mass<sub>HCl</sub> = 36.5 g/mol and molar mass<sub>NH<sub>3</sub></sub> = 17.0 g/mol.

The ratio of diffusion rates is 1.47.

#### 3 EVALUATE THE ANSWER

A ratio of roughly 1.5 is logical because molecules of ammonia are about half as massive as molecules of hydrogen chloride. Because the molar masses have three significant figures, the answer also does. Note that the units cancel, and the answer is stated correctly without any units.

### APPLICATIONS

1. Calculate the ratio of effusion rates for nitrogen (N<sub>2</sub>) and neon (Ne).
2. Calculate the ratio of diffusion rates for carbon monoxide and carbon dioxide.
3. **Challenge** 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?

### IN-CLASS Example

**Question** Calculate the ratio of effusion rates for helium and argon.

#### Answer

$$\begin{aligned} \frac{\text{Rate}_{\text{He}}}{\text{Rate}_{\text{Ar}}} &= \sqrt{\frac{\text{atomic mass}_{\text{Ar}}}{\text{atomic mass}_{\text{He}}}} \\ &= \sqrt{\frac{39.948 \text{ g/mol}}{4.003 \text{ g/mol}}} \\ &= 3.159 \end{aligned}$$

### APPLICATIONS

Have students refer to the Selected Solutions appendix for complete solutions to odd-numbered problems.

1.  $R_{\text{N}_2}/R_{\text{Ne}} = 0.8488$
2. 1.253
3. 2.5 mol/min

## Dalton's Law of Partial Pressures

$$P_{\text{total}} = P_1 + P_2 + P_3 + \dots P_n$$

$P_{\text{total}}$  represents total pressure.  $P_1$ ,  $P_2$ , and  $P_3$  represent the partial pressures of each gas up to the final gas,  $P_n$ .

To calculate the total pressure of a mixture of gases, add the partial pressures of each of the gases in the mixture.

### EXAMPLE 2

**THE PARTIAL PRESSURE OF A GAS** A mixture of oxygen ( $\text{O}_2$ ), carbon dioxide ( $\text{CO}_2$ ), and nitrogen ( $\text{N}_2$ ) has a total pressure of 0.97 atm. What is the partial pressure of  $\text{O}_2$  if the partial pressure of  $\text{CO}_2$  is 0.70 atm and the partial pressure of  $\text{N}_2$  is 0.12 atm?

#### 1 ANALYZE THE PROBLEM

You are given the total pressure of a mixture and the partial pressure of two gases in the mixture. To find the partial pressure of the third gas, use the equation that relates partial pressures to total pressure.

##### Known

$$P_{\text{N}_2} = 0.12 \text{ atm}$$

$$P_{\text{CO}_2} = 0.70 \text{ atm}$$

$$P_{\text{total}} = 0.97 \text{ atm}$$

##### Unknown

$$P_{\text{O}_2} = ? \text{ atm}$$

#### 2 SOLVE FOR THE UNKNOWN

$$P_{\text{total}} = P_{\text{N}_2} + P_{\text{CO}_2} + P_{\text{O}_2}$$

State Dalton's law of partial pressures.

$$P_{\text{O}_2} = P_{\text{total}} - P_{\text{CO}_2} - P_{\text{N}_2}$$

Solve for  $P_{\text{O}_2}$ .

$$P_{\text{O}_2} = 0.97 \text{ atm} - 0.70 \text{ atm} - 0.12 \text{ atm}$$

Substitute  $P_{\text{N}_2} = 0.12 \text{ atm}$ ,  $P_{\text{CO}_2} = 0.70 \text{ atm}$ , and  $P_{\text{total}} = 0.97 \text{ atm}$ .

$$P_{\text{O}_2} = 0.15 \text{ atm}$$

#### 3 EVALUATE THE ANSWER

Adding the calculated value for the partial pressure of oxygen to the known partial pressures gives the total pressure, 0.97 atm. The answer has two significant figures to match the data.

### APPLICATIONS

- What is the partial pressure of hydrogen gas in a mixture of hydrogen and helium if the total pressure is 600 mmHg and the partial pressure of helium is 439 mmHg?
- 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.
- 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.
- Challenge** Air is a mixture of gases. By percentage, it is roughly 78 percent nitrogen, 21 percent oxygen, and 1 percent argon. (There are trace amounts of many other gases in air.) If the atmospheric pressure is 760 mmHg, what are the partial pressures of nitrogen, oxygen, and argon in the atmosphere?

### IN-CLASS Example

**Question** A mixture of oxygen ( $\text{O}_2$ ), dinitrogen monoxide ( $\text{N}_2\text{O}$ ), and argon ( $\text{Ar}$ ) has a total pressure of 0.98 atm. What is the partial pressure of  $\text{N}_2\text{O}$ , if the partial pressure of  $\text{O}_2$  is 0.48 atm and the partial pressure of  $\text{Ar}$  is 0.15 atm?

#### Answer

$$P_{\text{total}} = P_{\text{O}_2} + P_{\text{N}_2\text{O}} + P_{\text{Ar}}$$

$$P_{\text{N}_2\text{O}} = P_{\text{total}} - P_{\text{O}_2} - P_{\text{Ar}}$$

$$P_{\text{N}_2\text{O}} = 0.98 \text{ atm} - 0.48 \text{ atm} - 0.15 \text{ atm}$$

$$P_{\text{N}_2\text{O}} = 0.35 \text{ atm}$$

### APPLICATIONS

Have students refer to the Selected Solutions appendix for complete solutions to odd-numbered problems.

4. 161 mmHg

5. 13.78 kPa

6. 10.2 kPa

7.  $\text{N}_2 = 590 \text{ mmHg}$ ;  $\text{O}_2 = 160 \text{ mmHg}$ ;  
 $\text{Ar} = 8 \text{ mmHg}$

## Section 1 Review

### SECTION 1 REVIEW

#### Section Summary

- The kinetic-molecular theory explains the properties of gases in terms of the size, motion, and energy of their particles.
- Dalton's law of partial pressures is used to determine the pressures of individual gases in gas mixtures.
- Graham's law is used to compare the diffusion rates of two gases.

- 8. MAIN IDEA Explain** Use the kinetic theory to explain the behavior of gases.
- 9. Describe** how the mass of a gas particle affects its rate of effusion and diffusion.
- 10. Explain** how gas pressure is measured.
- 11. Explain** why the container of water must be inverted when a gas is collected by displacement of water.
- 12. Calculate** 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?
- 13. Infer** whether or not temperature has any effect on the diffusion rate of a gas. Explain your answer.

### SECTION 1 REVIEW

8. Gases consist of small particles in random motion, which experience elastic collisions.
9. The rate of effusion and diffusion decreases as mass increases.
10. Atmospheric pressure is measured using a barometer. The gas pressure in a closed container is measured using a manometer.
11. 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.
12. 0.45 atm
13. As temperature increases, the velocity of the particles increase and the particles will diffuse faster.

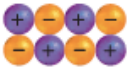

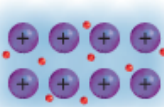
## Section 2 “Forces of Attraction”

### Student learning outcomes-

- Explain how the physical and chemical properties of a solid or liquid depend on the present particles, the type of bonds, and the intermolecular and intramolecular forces.

### New vocabulary-

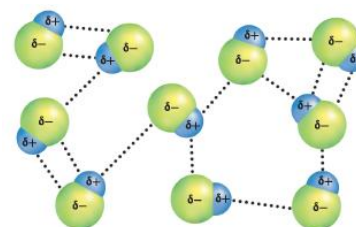
- **Polar covalent:** a type of bond that forms when electrons are not shared equally.

Force	Model	Basis of Attraction	Example
Ionic		cations and anions	NaCl
Covalent		positive nuclei and shared electrons	H <sub>2</sub>
Metallic		metal cations and mobile electrons	Fe

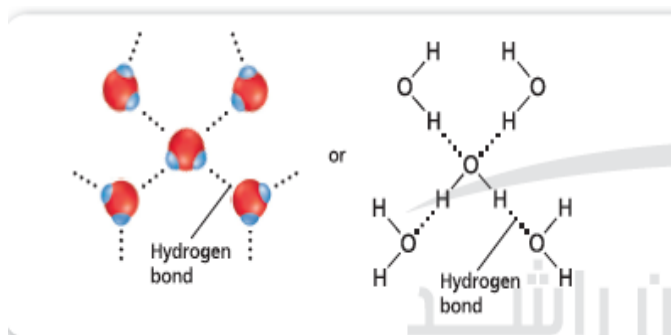
- **Dispersion force:** are weak forces that result from temporary shifts in the density of electrons in electron clouds.



- **Dipole-dipole force:** attractions between oppositely charged regions of polar molecules.



- **Hydrogen bond:** is a dipole-dipole attraction that occurs between molecules containing a hydrogen atom bonded to a small, highly electronegative atom with at least one lone electron pair.



## Section 2 Review

### SECTION 2 REVIEW

#### Section Summary

- Intramolecular forces are stronger than intermolecular forces.
- Dispersion forces are intermolecular forces between temporary dipoles.
- Dipole-dipole forces occur between polar molecules.

**14. MAIN IDEA Explain** what determines a substance's state at a given temperature.

**15. Compare and contrast** intermolecular forces and describe intramolecular forces.

**16. Evaluate** 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.  $H_2S$

c.  $HCl$

d.  $HF$

**17. Interpret Data** In a methane molecule ( $CH_4$ ), there are four 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?

### SECTION 2 REVIEW

- 14.** The intermolecular forces between the particles determine the state of a substance. In a solid, the intermolecular forces are very strong and hold the particles together. In a liquid, the intermolecular forces are weaker and in a gas, the particles no longer experience intermolecular forces.
- 15.** Intermolecular forces occur between particles. Intramolecular forces hold particles together.

- 16.** Hydrogen bonds: d; only dispersion forces: a; d is a polar molecule with an oxygen, nitrogen, or fluorine atom bonded to hydrogen. Molecule a is nonpolar.
- 17.** More bonds mean more electrons to form temporary dipoles, which means greater dispersion forces. Methane is a gas; octane is a liquid.

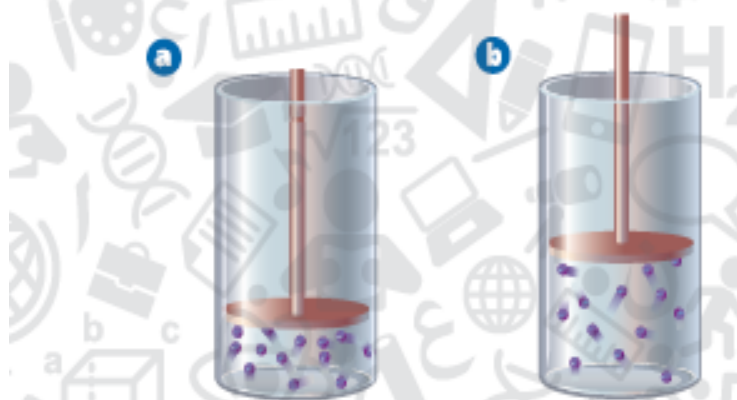


## Chapter 8 Review and MS

### SECTION 1

#### Mastering Concepts

- 34. What is an elastic collision?
- 35. How does the kinetic energy of particles vary as a function of temperature?
- 36. Use the kinetic-molecular theory to explain the compression and expansion of gases.
- 37. List the three basic assumptions of the kinetic-molecular theory.
- 38. Describe the common properties of gases.
- 39. Compare diffusion and effusion. Explain the relationship between the rates of these processes and the molar mass of a gas.



■ Figure 31

- 40. In Figure 31, what happens to the density of gas particles in the cylinder as the piston moves from Position A to Position B?
- 41. **Baking** 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 a shorter cooking time at a high elevation?

- 34. one in which no kinetic energy is lost
- 35. It is directly proportional to their temperature.
- 36. 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.
- 37. (1). Matter is composed of small particles.  
(2). The particles are in constant motion and undergo elastic collisions.  
(3). The particles have kinetic energy and the average kinetic energy of the particles is temperature.
- 38. Gases have low density, can be compressed, will expand to fill all available space, and can undergo diffusion and effusion.
- 39. Both involve the movement of gas particles. Diffusion is the movement of one substance through another; effusion is when a substance under pressure escapes through a tiny opening. Effusion and diffusion rates are inversely related to molecular mass of a gas.
- 40. Density decreases because the gas particles occupy more volume per unit mass.
- 41. 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.

## Mastering Problems

42. What is the molar mass of a gas that takes three times longer to effuse than helium?
43. What is the ratio of effusion rates of krypton and neon at the same temperature and pressure?
44. Calculate the molar mass of a gas that diffuses three times faster than oxygen under similar conditions.
45. 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?
46. 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?
47. **Mountain Climbing** 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?
48. **High Altitude** The atmospheric pressure in Denver, Colorado, is usually about 84.0 kPa. What is this pressure in atm and torr units?
49. At an ocean depth of 76.2 m, the pressure is about 8.4 atm. Convert the pressure to mmHg and kPa units.

42. 36.0 g/mol
43.  $\text{Rate}_{\text{Kr}}/\text{Rate}_{\text{Ne}} = 0.4907$
44. 3.56 g/mol
45. 0.01 atm
46. 0.99 atm
47. 0.332 atm; It is about one-third of the 1-atm pressure at sea level.
48. 84.0 kPa = 0.829 atm and  $6.30 \times 10^2$  torr
49. 8.4 atm =  $8.5 \times 10^2$  kPa and  $6.4 \times 10^3$  mmHg
50. The gases will diffuse until both bulbs are filled with the same gas mixtures.

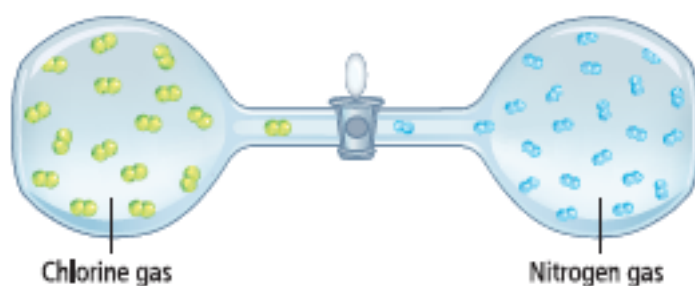


Figure 32

50. **Figure 32** represents an experimental set-up in which the left bulb is filled with chlorine gas and the right bulb is filled with nitrogen gas. Describe what happens when the stopcock is opened. Assume that the temperature of the system is held constant during the experiment.



## SECTION 2

### Mastering Concepts

51. Explain the difference between a temporary dipole and a permanent dipole.
52. Why are dispersion forces weaker than dipole-dipole forces?
53. Explain why hydrogen bonds are stronger than most dipole-dipole forces.
54. Compare intramolecular and intermolecular forces.
55. Hypothesize why long, nonpolar molecules would interact more strongly with one another than spherical nonpolar molecules of similar composition.

### Mastering Problems

56. **Polar Molecules** Use relative differences in electronegativity to label the ends of the polar molecules listed as partially positive or partially negative.  
a. HF      b. HBr      c. NO      d. CO
57. Draw the structure of the dipole-dipole interaction between two molecules of carbon monoxide.
58. Decide which of the substances listed can form hydrogen bonds.  
a.  $\text{H}_2\text{O}$       b.  $\text{H}_2\text{O}_2$       c. HF      d.  $\text{NH}_3$
59. Decide which one of the molecules listed below can form intermolecular hydrogen bonds, and then draw it, showing several molecules attached together by hydrogen bonds.  
a. NaCl      b.  $\text{MgCl}_2$       c.  $\text{H}_2\text{O}_2$       d.  $\text{CO}_2$

51. A temporary dipole forms when one molecule is close to another molecule, and the electrons repel each other creating a greater electron density in one part of the molecule. Permanent dipoles are found in polar molecules in which some regions of the molecule are always partially positive and partially negative.
52. Dispersion forces are between temporary dipoles. Dipole-dipole forces are between permanent dipoles.

53. 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.
54. Intramolecular forces hold atoms together in a molecule while intermolecular forces hold different molecules together.
55. Because long molecules have greater surface areas, more intermolecular forces can exist.

56. a.  $\text{H}^+ - \text{F}^-$ ;    b.  $\text{H}^+ - \text{Br}^-$ ;    c.  $\text{N}^+ - \text{O}^-$ ;    d.  $\text{C}^+ - \text{O}^-$
57. Refer to Figure 9. The drawing should show two CO molecules, with the C partially positive and the O partially negative. The C of each molecule should be bonded to the O of the other.
58. All of the substances can form hydrogen bonds.
59.  $\text{H}_2\text{O}_2$  can form hydrogen bonds. Refer to the Solutions Manual for drawing.

## Chapter 9 Gases

### Section 1" The Gas Laws"

#### Student learning outcomes-

- Use the kinetic molecular theory to explain the properties and behaviour of gases.
- Employ gas laws to describe the behaviour of gases and their practical applications.

#### New vocabulary-

- **Scientific law:** describes a relationship in nature that is supported by many experiments.
- **Boyle's law:** states that the volume of a fixed amount of gas held at a constant temperature varies inversely with the pressure.

#### Boyle's Law

$$P_1 V_1 = P_2 V_2 \quad P \text{ represents pressure. } V \text{ represents volume.}$$

For a given amount of gas held at constant temperature, the product of pressure and volume is a constant.

$P_1$  and  $V_1$  represent the initial conditions, and  $P_2$  and  $V_2$  represent new conditions. If you know any three of these values, you can solve for the fourth by rearranging the equation.

- **Absolute zero:** Zero on the Kelvin scale.
- **Charles's law:** states that the volume of a given amount of gas is directly proportional to its Kelvin temperature at constant pressure.

#### Charles's Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \begin{array}{l} V \text{ represents volume.} \\ T \text{ represents temperature.} \end{array}$$

In the equation above,  $V_1$  and  $T_1$  represent initial conditions, while  $V_2$  and  $T_2$  are new conditions. As with Boyle's law, if you know three of the values, you can calculate the fourth.

The temperature must be expressed in kelvins when using the equation for Charles's law. To convert a temperature from Celsius degrees to kelvins, add 273 to the Celsius temperature:

$$T_k = 273 + T_c$$

- **Gay-Lussac's law:** states that the pressure of a fixed amount of gas varies directly with the Kelvin temperature when the volume remains constant.

### Gay-Lussac's Law

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

*P* represents pressure.  
*T* represents temperature.

For a given amount of gas held at constant volume, the quotient of the pressure and the Kelvin temperature is a constant.

- **Combined gas law** states the relationships between pressure, temperature, and volume of a fixed amount of gas.

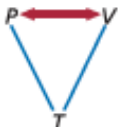
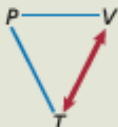
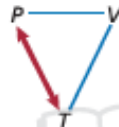

### The Combined Gas Law

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

*P* represents pressure. *V* represents volume.  
*T* represents temperature.

For a given amount of gas, the product of pressure and volume, divided by the Kelvin temperature, is a constant.

## The Gas Laws -

Table 1 The Gas Laws				
Law	Boyle's	Charles's	Gay-Lussac's	Combined
Formula	$P_1 V_1 = P_2 V_2$	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$	$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$
What is constant?	amount of gas, temperature	amount of gas, pressure	amount of gas, volume	amount of gas
Graphic organizer				

## Worked examples –

### EXAMPLE 1

**BOYLE'S LAW** A diver blows a 0.75 L air bubble 10 m under water. As it rises to the surface, the pressure goes from 2.25 atm to 1.03 atm. What will be the volume of air in the bubble at the surface?

#### 1 ANALYZE THE PROBLEM

According to Boyle's law, the decrease in pressure on the bubble will result in an increase in volume, so the initial volume should be multiplied by a pressure ratio greater than 1.

**Known**

$$V_1 = 0.75 \text{ L}$$

$$P_1 = 2.25 \text{ atm}$$

$$P_2 = 1.03 \text{ atm}$$

**Unknown**

$$V_2 = ? \text{ L}$$

#### 2 SOLVE FOR THE UNKNOWN

Use Boyle's law. Solve for  $V_2$ , and calculate the new volume.

$$P_1 V_1 = P_2 V_2$$

State Boyle's law.

$$V_2 = V_1 \left( \frac{P_1}{P_2} \right)$$

Solve for  $V_2$ .

$$V_2 = 0.75 \text{ L} \left( \frac{2.25 \text{ atm}}{1.03 \text{ atm}} \right)$$

Substitute  $V_1 = 0.75 \text{ L}$ ,  $P_1 = 2.25 \text{ atm}$ , and  $P_2 = 1.03 \text{ atm}$ .

$$V_2 = 0.75 \text{ L} \left( \frac{2.25 \text{ atm}}{1.03 \text{ atm}} \right) = 1.6 \text{ L}$$

Multiply and divide numbers and units.

#### 3 EVALUATE THE ANSWER

The pressure decreases by roughly half, so the volume should roughly double. The answer is expressed in liters, a unit of volume, and correctly contains two significant figures.

### APPLICATIONS

Assume that the temperature and the amount of gas are constant in the following problems.

1. The volume of a gas at 99.0 kPa is 300.0 mL. If the pressure is increased to 188 kPa, what will be the new volume?
2. The pressure of a sample of helium in a 1.00 L container is 0.988 atm. What is the new pressure if the sample is placed in a 2.00 L container?
3. **Challenge** Air trapped in a cylinder fitted with a piston occupies 145.7 mL at 1.08 atm pressure. What is the new volume when the piston is depressed, increasing the pressure by 25%?

## Applications MS –

1. 158 mL .
2. 0.494 atm
3. 117 mL

## EXAMPLE 2

**CHARLES'S LAW** A helium balloon in a closed car occupies a volume of 2.32 L at 40.0°C. If the car is parked on a hot day and the temperature inside rises to 75.0°C, what is the new volume of the balloon, assuming the pressure remains constant?

### 1 ANALYZE THE PROBLEM

Charles's law states that as the temperature of a fixed amount of gas increases, so does its volume, assuming constant pressure. Therefore, the volume of the balloon will increase. The initial volume should be multiplied by a temperature ratio greater than 1.

#### Known

$$T_1 = 40.0^\circ\text{C}$$

$$V_1 = 2.32 \text{ L}$$

$$T_2 = 75.0^\circ\text{C}$$

#### Unknown

$$V_2 = ? \text{ L}$$

### 2 SOLVE FOR THE UNKNOWN

Convert degrees Celsius to kelvins.

$$T_K = 273 + T_C$$

Apply the conversion factor.

$$T_1 = 273 + 40.0^\circ\text{C} = 313.0 \text{ K}$$

Substitute  $T_1 = 40.0^\circ\text{C}$ .

$$T_2 = 273 + 75.0^\circ\text{C} = 348.0 \text{ K}$$

Substitute  $T_2 = 75.0^\circ\text{C}$ .

Use Charles's law. Solve for  $V_2$ , and substitute the known values into the rearranged equation.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

State Charles's law.

$$V_2 = V_1 \left( \frac{T_2}{T_1} \right)$$

Solve for  $V_2$ .

$$V_2 = 2.32 \text{ L} \left( \frac{348.0 \text{ K}}{313.0 \text{ K}} \right)$$

Substitute  $V_1 = 2.32 \text{ L}$ ,  $T_1 = 313.0 \text{ K}$ , and  $T_2 = 348.0 \text{ K}$ .

$$V_2 = 2.32 \text{ L} \left( \frac{348.0 \cancel{\text{K}}}{313.0 \cancel{\text{K}}} \right) = 2.58 \text{ L}$$

Multiply and divide numbers and units.

### 3 EVALUATE THE ANSWER

The increase in kelvins is relatively small, so the volume should show a small increase. The unit of the answer is liters, a volume unit, and there are three significant figures.

## APPLICATIONS

Assume that the pressure and the amount of gas remain constant in the following problems.

- What volume will the gas in the balloon at right occupy at 250 K?
- A gas at 89°C occupies a volume of 0.67 L. At what Celsius temperature will the volume increase to 1.12 L?
- The Celsius temperature of a 3.00-L sample of gas is lowered from 80.0°C to 30.0°C. What will be the resulting volume of this gas?
- Challenge** A gas occupies 0.67 L at 350 K. What temperature is required to reduce the volume by 45%?



- 3.1 L
- 330°C
- 2.58 L
- 190 K



### EXAMPLE 3

**GAY-LUSSAC'S LAW** The pressure of the oxygen gas inside a canister is 5.00 atm at 25.0°C. The canister is located at a camp high on Jais in UAE. If the temperature there falls to -10.0°C, what is the new pressure inside the canister?

#### 1 ANALYZE THE PROBLEM

Gay-Lussac's law states that if the temperature of a gas decreases, so does its pressure when volume is constant. Therefore, the pressure in the oxygen canister will decrease. The initial pressure should be multiplied by a temperature ratio less than 1.

##### Known

$$P_1 = 5.00 \text{ atm}$$

$$T_1 = 25.0^\circ\text{C}$$

$$T_2 = -10.0^\circ\text{C}$$

##### Unknown

$$P_2 = ? \text{ atm}$$

#### 2 SOLVE FOR THE UNKNOWN

Convert degrees Celsius to kelvins.

$$T_K = 273 + T_C$$

Apply the conversion factor.

$$T_1 = 273 + 25.0^\circ\text{C} = 298.0 \text{ K}$$

Substitute  $T_1 = 25.0^\circ\text{C}$ .

$$T_2 = 273 + (-10.0^\circ\text{C}) = 263.0 \text{ K}$$

Substitute  $T_2 = -10.0^\circ\text{C}$ .

Use Gay-Lussac's law. Solve for  $P_2$ , and substitute the known values into the rearranged equation.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

State Gay-Lussac's law.

$$P_2 = P_1 \left( \frac{T_2}{T_1} \right)$$

Solve for  $P_2$ .

$$P_2 = 5.00 \text{ atm} \left( \frac{263.0 \text{ K}}{298.0 \text{ K}} \right)$$

Substitute  $P_1 = 5.00 \text{ atm}$ ,  
 $T_1 = 298.0 \text{ K}$ , and  $T_2 = 263.0 \text{ K}$ .

$$P_2 = 5.00 \text{ atm} \left( \frac{263.0 \text{ K}}{298.0 \text{ K}} \right) = 4.41 \text{ atm}$$

Multiply and divide numbers  
and units.

#### 3 EVALUATE THE ANSWER

Kelvin temperature decreases, so the pressure should decrease. The unit is atm, a pressure unit, and there are three significant figures.

### APPLICATIONS

Assume that the volume and the amount of gas are constant in the following problems.

- The pressure in an automobile tire is 1.88 atm at 25.0°C. What will be the pressure if the temperature increases to 37.0°C?
- Helium gas in a 2.00 L cylinder is under 1.12 atm pressure. At 36.5°C, that same gas sample has a pressure of 2.56 atm. What was the initial temperature in degrees Celsius of the gas in the cylinder?
- Challenge** If a gas sample has a pressure of 30.7 kPa at 0.00°C, by how many degrees Celsius does the temperature have to increase to cause the pressure to double?

8. 1.96 atm

9. -138°C

10. 273°C

## EXAMPLE 4

**THE COMBINED GAS LAW** A gas at 110 kPa and 30.0°C fills a flexible container with an initial volume of 2.00 L. If the temperature is raised to 80.0°C and the pressure increases to 440 kPa, what is the new volume?

### 1 ANALYZE THE PROBLEM

Both pressure and temperature change, so you will need to use the combined gas law. The pressure quadruples, but the temperature does not increase by such a large factor. Therefore, the new volume will be smaller than the starting volume.

#### Known

$$\begin{aligned} P_1 &= 110 \text{ kPa} & P_2 &= 440 \text{ kPa} \\ T_1 &= 30.0^\circ\text{C} & T_2 &= 80.0^\circ\text{C} \\ V_1 &= 2.00 \text{ L} \end{aligned}$$

#### Unknown

$$V_2 = ? \text{ L}$$

### 2 SOLVE FOR THE UNKNOWN

Convert degrees Celsius to kelvins.

$$T_K = 273 + T_C$$

$$T_1 = 273 + 30.0^\circ\text{C} = 303.0 \text{ K}$$

$$T_2 = 273 + 80.0^\circ\text{C} = 353.0 \text{ K}$$

Apply the conversion factor.

Substitute  $T_1 = 30.0^\circ\text{C}$ .

Substitute  $T_2 = 80.0^\circ\text{C}$ .

Use the combined gas law. Solve for  $V_2$ , and substitute the known values into the rearranged equation.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

State the combined gas law.

$$V_2 = V_1 \left( \frac{P_1}{P_2} \right) \left( \frac{T_2}{T_1} \right)$$

Solve for  $V_2$ .

$$V_2 = 2.00 \text{ L} \left( \frac{110 \text{ kPa}}{440 \text{ kPa}} \right) \left( \frac{353.0 \text{ K}}{303.0 \text{ K}} \right)$$

Substitute  $V_1 = 2.00 \text{ L}$ ,  $P_1 = 110 \text{ kPa}$ ,  $P_2 = 440 \text{ kPa}$ ,  $T_2 = 353.0 \text{ K}$ , and  $T_1 = 303.0 \text{ K}$ .

$$V_2 = 2.00 \text{ L} \left( \frac{110 \text{ kPa}}{440 \text{ kPa}} \right) \left( \frac{353.0 \text{ K}}{303.0 \text{ K}} \right) = 0.58 \text{ L}$$

Multiply and divide numbers and units.

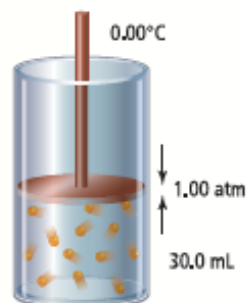
### 3 EVALUATE THE ANSWER

Because the pressure change is much greater than the temperature change, the volume undergoes a net decrease. The unit is liters, a volume unit, and there are two significant figures.

## APPLICATIONS

Assume that the amount of gas is constant in the following problems.

11. A sample of air in a syringe exerts a pressure of 1.02 atm at 22.0°C. The syringe is placed in a boiling-water bath at 100.0°C. The pressure is increased to 1.23 atm by pushing the plunger in, which reduces the volume to 0.224 mL. What was the initial volume?
12. A balloon contains 146.0 mL of gas confined at a pressure of 1.30 atm and a temperature of 5.0°C. If the pressure doubles and the temperature decreases to 2.0°C, what will be the volume of gas in the balloon?
13. **Challenge** If the temperature in the gas cylinder at right increases to 30.0°C and the pressure increases to 1.20 atm, will the cylinder's piston move up or down?



11. 0.214 mL
12. 72 mL
13. down

## Section 1 Review

### SECTION 1 REVIEW

#### Section Summary

- Boyle's law states that the volume of a fixed amount of gas is inversely proportional to its pressure at constant temperature.
- Charles's law states that the volume of a fixed amount of gas is directly proportional to its Kelvin temperature at constant pressure.
- Gay-Lussac's law states that the pressure of a fixed amount of gas is directly proportional to its Kelvin temperature at constant volume.
- The combined gas law relates pressure, temperature, and volume in a single statement.

**14. MAIN IDEA State** the relationships between pressure, temperature, and volume of a fixed amount of gas.

**15. Explain** Which of the three variables that apply to equal amounts of gases are directly proportional? Which are inversely proportional?

**16. Analyze** A weather balloon is released into the atmosphere. You know the initial volume, temperature, and air pressure. What information will you need to predict its volume when it reaches its final altitude? Which law would you use to calculate this volume?

**17. Infer** why gases such as the oxygen used at hospitals are compressed. Why must compressed gases be shielded from high temperatures? What must happen to compressed oxygen before it can be inhaled?

**18. Calculate** A rigid plastic container holds 1.00 L of methane gas at 660 torr pressure when the temperature is 22.0°C. How much pressure will the gas exert if the temperature is raised to 44.6°C?

**19. Design a concept map** that shows the relationships between pressure, volume, and temperature in Boyle's, Charles's, and Gay-Lussac's laws.

### SECTION 1 REVIEW

**14.** This relationship is given by the combined gas law:  $P_1V_1/T_1 = P_2V_2/T_2$ . For example: when the temperature increases, either the volume or pressure increases (or both).

**15.**  $P$  and  $V$  are directly proportional to  $T$ , and  $P$  and  $V$  are inversely proportional to each other.

**16.** You would need to know the final temperature and final pressure to calculate the final volume. Use the combined gas law.

**17.** A greater mass confined to a smaller volume makes transporting and storing of gases easier. Increasing temperature increases pressure, and the cylinders might explode. Before compressed oxygen can be breathed, it must be decompressed.

**18.** 711 torr

**19.** The concept map should show how  $P$ ,  $V$ , and  $T$  are proportional to one another. Students should also label each pair of variables used in the gas laws. Refer to the Solutions Manual.

## Section 2 “The Ideal Gas Law”

### Student learning outcomes-

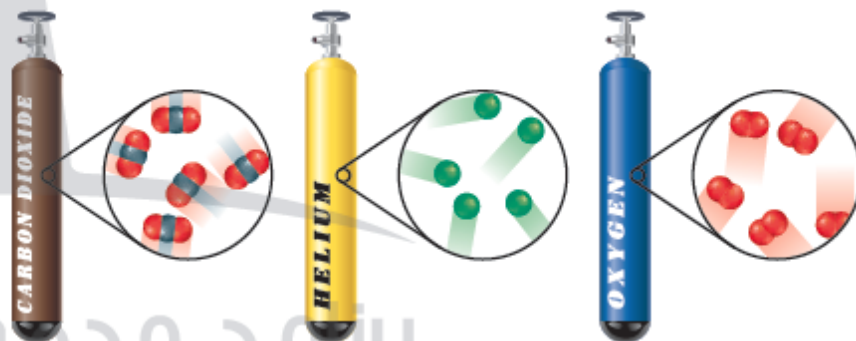
- Use the kinetic molecular theory to explain the properties and behaviour of gases.
- Use the kinetic molecular theory to explain the properties and behaviour of gases.
- Apply Avogadro's hypothesis in calculating gas masses and volumes.

### New vocabulary-

- **Mole:** an SI base unit used to measure the amount of a substance; the amount of a pure substance that contains  $6.02 \times 10^{23}$  representative particles.
- **Avogadro's Principle:** states that equal volumes of gases at the same temperature and pressure contain equal numbers of particles.

■ **Figure 5** Gas tanks of equal volume that are at the same pressure and temperature contain equal numbers of gas particles, regardless of which gas they contain.

**Infer** Why doesn't Avogadro's principle apply to liquids and solids?



- **Molar Volume (of a gas):** is the volume that 1 mol occupies at 0.00 Celsius and 1.00 atm pressure.
- **Standard temperature and pressure (STP):** is the condition of 0.00 Celsius and 1.00 atm pressure.
- **Ideal Gas Constant:** Experiments using known values of P, T, V, and n have determined the value of this constant. It is called the ideal gas constant, and it is represented by the symbol R.

Table 2 Values of R	
Value of R	Units of R
0.0821	$\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$
8.314	$\frac{\text{L}\cdot\text{kPa}}{\text{mol}\cdot\text{K}}$
62.4	$\frac{\text{L}\cdot\text{mm Hg}}{\text{mol}\cdot\text{K}}$

- **Ideal Gas Law:** describes the physical behavior of an ideal gas in terms of the pressure, volume, temperature, and number of moles of gas present.

### The Ideal Gas Law

$$PV = nRT$$

$P$  represents pressure.  $V$  represents volume.

$n$  represents number of moles.  $R$  is the ideal gas constant.

$T$  represents temperature.

For a given amount of gas held at constant temperature, the product of pressure and volume is a constant.

## Worked examples –

### EXAMPLE 5

**MOLAR VOLUME** The main component of natural gas used for home heating and cooking is methane ( $\text{CH}_4$ ). Calculate the volume that 2.00 kg of methane gas will occupy at STP.

#### 1 ANALYZE THE PROBLEM

The number of moles can be calculated by dividing the mass of the sample,  $m$ , by its molar mass,  $M$ . The gas is at STP (0.00°C and 1.00 atm pressure), so you can use the molar volume to convert from the number of moles to the volume.

**Known**

$$m = 2.00 \text{ kg}$$

$$T = 0.00^\circ\text{C}$$

$$P = 1.00 \text{ atm}$$

**Unknown**

$$V = ? \text{ L}$$

#### 2 SOLVE FOR THE UNKNOWN

Determine the molar mass for methane.

$$M = 1 \text{ C atom} \left( \frac{12.01 \text{ amu}}{1 \text{ C atom}} \right) + 4 \text{ H atoms} \left( \frac{1.01 \text{ amu}}{1 \text{ H atom}} \right)$$

$$= 12.01 \text{ amu} + 4.04 \text{ amu} = 16.05 \text{ amu}$$

$$= 16.05 \text{ g/mol}$$

Determine the number of moles of methane.

$$2.00 \text{ kg} \left( \frac{1000 \text{ g}}{1 \text{ kg}} \right) = 2.00 \times 10^3 \text{ g}$$

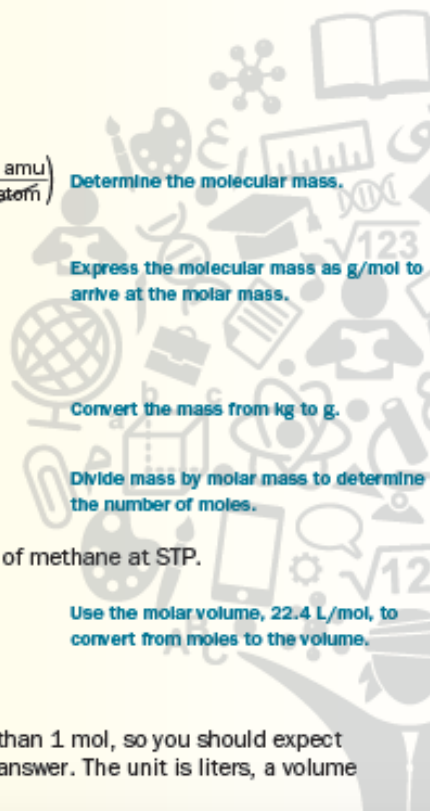
$$\frac{m}{M} = \frac{2.00 \times 10^3 \text{ g}}{16.05 \text{ g/mol}} = 125 \text{ mol}$$

Use the molar volume to determine the volume of methane at STP.

$$V = 125 \text{ mol} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 2.80 \times 10^3 \text{ L}$$

#### 3 EVALUATE THE ANSWER

The amount of methane present is much more than 1 mol, so you should expect a large volume, which is in agreement with the answer. The unit is liters, a volume unit, and there are three significant figures.





## APPLICATIONS

20. What size container do you need to hold 0.0459 mol of  $\text{N}_2$  gas at STP?
21. How much carbon dioxide gas, in grams, is in a 1.0 L balloon at STP?
22. What volume in milliliters will 0.00922 g of  $\text{H}_2$  gas occupy at STP?
23. What volume will 0.416 g of krypton gas occupy at STP?
24. Calculate the volume that 4.5 kg of ethylene gas ( $\text{C}_2\text{H}_4$ ) will occupy at STP.
25. **Challenge** A flexible plastic container contains 0.860 g of helium gas in a volume of 19.2 L. If 0.205 g of helium is removed at constant pressure and temperature, what will be the new volume?

20. 1.03 L
21. 2.0 g
22. 102 mL
23. 0.111 L
24.  $3.6 \times 10^3$  L
25. 14.6 L

## EXAMPLE 6

**THE IDEAL GAS LAW** Calculate the number of moles of ammonia gas ( $\text{NH}_3$ ) contained in a 3.0 L vessel at  $3.00 \times 10^2$  K with a pressure of 1.50 atm.

### 1 ANALYZE THE PROBLEM

You are given the volume, temperature, and pressure of a gas sample. Use the ideal gas law, and select the value of  $R$  that contains the pressure units given in the problem. Because the pressure and temperature are close to STP, but the volume is much smaller than 22.4 L, it would make sense if the calculated answer were much smaller than 1 mol.

#### Known

$$\begin{aligned} V &= 3.0 \text{ L} \\ T &= 3.00 \times 10^2 \text{ K} \\ P &= 1.50 \text{ atm} \\ R &= 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \end{aligned}$$

#### Unknown

$$n = ? \text{ mol}$$

### 2 SOLVE FOR THE UNKNOWN

Use the ideal gas law. Solve for  $n$ , and substitute the known values.

$$PV = nRT$$

$$n = \frac{PV}{RT}$$

$$n = \frac{(1.50 \text{ atm})(3.0 \text{ L})}{\left(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(3.00 \times 10^2 \text{ K})}$$

$$n = \frac{(1.50 \text{ atm})(3.0 \text{ L})}{\left(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(3.00 \times 10^2 \text{ K})} = 0.18 \text{ mol}$$

State the Ideal gas law.

Solve for  $n$ .

Substitute  $V = 3.0 \text{ L}$ ,  
 $T = 3.00 \times 10^2 \text{ K}$ ,  
 $P = 1.50 \text{ atm}$ , and  $R =$   
 $0.0821 \text{ L} \cdot \text{atm}/\text{mol} \cdot \text{K}$ .

Multiply and divide  
numbers and units.

### 3 EVALUATE THE ANSWER

The answer agrees with the prediction that the number of moles present will be significantly less than 1 mol. The unit of the answer is the mole, and there are two significant figures.

## APPLICATIONS

26. Determine the Celsius temperature of 2.49 mol of a gas contained in a 1.00 L vessel at a pressure of 143 kPa.
27. Calculate the volume of a 0.323 mol sample of a gas at 265 K and 0.900 atm.
28. What is the pressure, in atmospheres, of a 0.108 mol sample of helium gas at a temperature of 20.0°C if its volume is 0.505 L?
29. If the pressure exerted by a gas at 25°C in a volume of 0.044 L is 3.81 atm, how many moles of gas are present?
30. **Challenge** An ideal gas has a volume of 3.0 L. If the number of moles of gas and the temperature are doubled, while the pressure remains constant, what is the new volume?

26.  $-266^{\circ}\text{C}$

27. 7.81 L

28. 5.14 atm

29.  $6.9 \times 10^{-3}$  mol

30. 12 L

## The Ideal Gas Law— Molar Mass and Density

The ideal gas law can be used to solve for the value of any one of the four variables  $P$ ,  $V$ ,  $T$ , or  $n$  if the values of the other three are known. However, you can also rearrange the  $PV = nRT$  equation to calculate the molar mass and density of a gas sample.

**Molar mass and the ideal gas law** To find the molar mass of a gas sample, the mass, temperature, pressure, and volume of the gas must be known. Recall that the number of moles of a gas ( $n$ ) is equal to the mass ( $m$ ) divided by the molar mass ( $M$ ). Therefore, the  $n$  in the equation can be replaced by  $m/M$ .

$$PV = nRT \quad \text{substitute } n = \frac{m}{M} \quad PV = \frac{mRT}{M}$$

You can rearrange the new equation to solve for the molar mass.

$$M = \frac{mRT}{PV}$$

**Density and the ideal gas law** Recall that the density ( $D$ ) of a substance is defined as mass ( $m$ ) per unit volume ( $V$ ). After rearranging the ideal gas equation to solve for molar mass, you can substitute  $D$  for  $m/V$ .

$$M = \frac{mRT}{pV} \quad \text{substitute } \frac{m}{V} = D \quad M = \frac{DRT}{p}$$

You can rearrange the new equation to solve for density.

$$D = \frac{MP}{RT}$$

## PROBLEM-SOLVING STRATEGIES

### Deriving Gas Laws

If you master the following strategy, you will need to remember only one gas law—the ideal gas law. Consider the example of a fixed amount of gas held at constant pressure. You need Charles's law to solve problems involving volume and temperature.

1. Use the ideal gas law to write two equations that describe the gas sample at two different volumes and temperatures. (Quantities that do not change are shown in **red**.)
2. Isolate volume and temperature—the two conditions that vary—on the same side of each equation.
3. Because  $n$ ,  $R$ , and  $P$  are constant under these conditions, you can set the volume and temperature conditions equal, deriving Charles's law.

$$\begin{array}{ccc} PV_1 = nRT_1 & & PV_2 = nRT_2 \\ \downarrow & & \downarrow \\ \frac{V_1}{T_1} = \frac{nR}{P} & & \frac{V_2}{T_2} = \frac{nR}{P} \\ \swarrow & & \searrow \\ \frac{V_1}{T_1} = \frac{V_2}{T_2} \end{array}$$

### Apply the Strategy

**Derive** Boyle's law, Gay-Lussac's law, and the combined gas law based on the example above.

Students should use the strategy to show the derivation from the ideal gas law to Boyle's law ( $P_1V_1 = P_2V_2$ ), Gay-Lussac's law ( $P_1/T_1 = P_2/T_2$ ), and the combined gas law ( $P_1V_1/T_1 = P_2V_2/T_2$ ).

## Section 2 Review

### SECTION 2 REVIEW

#### Section Summary

- Avogadro's principle states that equal volumes of gases at the same pressure and temperature contain equal numbers of particles.
- The ideal gas law relates the amount of a gas present to its pressure, temperature, and volume.
- The ideal gas law can be used to find molar mass if the mass of the gas is known or the density of the gas if its molar mass is known.
- At very high pressures and very low temperatures, real gases behave differently than ideal gases.

**31. MAIN IDEA Explain** why Avogadro's principle holds true for ideal gases that have small particles and for ideal gases that have large particles.

**32. State** the equation for the ideal gas law.

**33. Analyze** how the ideal gas law applies to real gases using the kinetic-molecular theory.

**34. Predict** the conditions under which a real gas might deviate from ideal behavior.

**35. List** common units for each variable in the ideal gas law.

**36. Calculate** A 2.00 L flask is filled with propane gas ( $\text{C}_3\text{H}_8$ ) at a pressure of 1.00 atm and a temperature of  $-15.0^\circ\text{C}$ . What is the mass of the propane in the flask?

**37. Make and Use Graphs** For every  $6^\circ\text{C}$  drop in temperature, the air pressure in a car's tires goes down by about 1 psi (14.7 psi = 1.00 atm). Make a graph illustrating the change in tire pressure from  $20^\circ\text{C}$  to  $-20^\circ\text{C}$  (assume 30.0 psi at  $20^\circ\text{C}$ ).

### SECTION 2 REVIEW

**31.** The size of any gas particle is so small compared to the volume of the gas, it is assumed that no particle has any volume of its own.

**32.**  $PV = nRT$

**33.** A real gas behaves most like an ideal gas under conditions that increase the distance and reduce the attractions among gas particles. The best conditions for that are high temperature and low pressure.

**34.** A real gas might deviate from ideal behavior under conditions that decrease the distance and increase the attractions among gas particles, such as low temperature and high pressure.

**35.**  $P$ : atm, mm Hg, torr, kPa;  $V$ : L, mL;  $T$ : K;  $n$ : mol

**36.** 4.16 g

**37.** Graph should show air pressure graphed with relation to temperature; the resulting plot will be a straight line showing a direct relationship between the variables.

## Section 3 “Gas Stoichiometry”

### Student learning outcomes-

- Employ gas laws to describe the behaviour of gases and their practical applications.
- Apply Avogadro's hypothesis in calculating gas masses and volumes
- Explain Avogadro's hypothesis and its impact in understanding the chemical reactions of gases.

### New vocabulary-

- **Coefficient:** the number written in front of a reactant or product in a chemical equation, which tells the smallest number of particles of the substance involved in the reaction.

### Worked Examples –

#### EXAMPLE 7

**VOLUME–VOLUME PROBLEMS** What volume of oxygen gas is needed for the complete combustion of 4.00 L of propane gas ( $\text{C}_3\text{H}_8$ )? Assume that pressure and temperature remain constant.

#### 1 ANALYZE THE PROBLEM

You are given the volume of a gaseous reactant in a chemical reaction. Remember that the coefficients in a balanced chemical equation provide the volume relationships of gaseous reactants and products.

**Known**

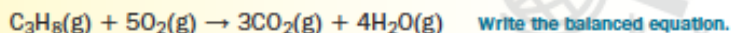
$$V_{\text{C}_3\text{H}_8} = 4.00 \text{ L}$$

**Unknown**

$$V_{\text{O}_2} = ? \text{ L}$$

#### 2 SOLVE FOR THE UNKNOWN

Use the balanced equation for the combustion of  $\text{C}_3\text{H}_8$ . Find the volume ratio for  $\text{O}_2$  and  $\text{C}_3\text{H}_8$ , then solve for  $V_{\text{O}_2}$ .



$$\frac{5 \text{ volumes O}_2}{1 \text{ volume C}_3\text{H}_8}$$

Find the volume ratio for  $\text{O}_2$  and  $\text{C}_3\text{H}_8$ .

$$\begin{aligned} V_{\text{O}_2} &= (4.00 \text{ L C}_3\text{H}_8) \times \frac{5 \text{ volumes O}_2}{1 \text{ volume C}_3\text{H}_8} \\ &= 20.0 \text{ L O}_2 \end{aligned}$$

Multiply the known volume of  $\text{C}_3\text{H}_8$  by the volume ratio to find the volume of  $\text{O}_2$ .

#### 3 EVALUATE THE ANSWER

The coefficients in the combustion equation show that a much larger volume of  $\text{O}_2$  than  $\text{C}_3\text{H}_8$  is used up in the reaction, which is in agreement with the calculated answer. The unit of the answer is liters, a unit of volume, and there are three significant figures.



## APPLICATIONS

38. How many liters of propane gas ( $\text{C}_3\text{H}_8$ ) will undergo complete combustion with 34.0 L of oxygen gas?
39. Determine the volume of hydrogen gas needed to react completely with 5.00 L of oxygen gas to form water.
40. What volume of oxygen is needed to completely combust 2.36 L of methane gas ( $\text{CH}_4$ )?
41. **Challenge** Nitrogen and oxygen gases react to form dinitrogen monoxide gas ( $\text{N}_2\text{O}$ ). What volume of  $\text{O}_2$  is needed to produce 34 L of  $\text{N}_2\text{O}$ ?

38. 6.80 L  $\text{C}_3\text{H}_8$

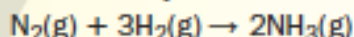
39. 10.0 L  $\text{H}_2$

40. 4.72 L  $\text{O}_2$

41. 17 L  $\text{O}_2$

## EXAMPLE 8

**VOLUME–MASS PROBLEMS** Ammonia is synthesized from hydrogen and nitrogen.



If 5.00 L of nitrogen reacts completely with hydrogen at a pressure of 3.00 atm and a temperature of 298 K, how much ammonia, in grams, is produced?

### 1 ANALYZE THE PROBLEM

You are given the volume, pressure, and temperature of a gas sample. The mole and volume ratios of gaseous reactants and products are given by the coefficients in the balanced chemical equation. Volume can be converted to moles and thus related to mass by using molar mass and the ideal gas law.

**Known**

$$V_{\text{N}_2} = 5.00 \text{ L}$$

$$P = 3.00 \text{ atm}$$

$$T = 298 \text{ K}$$

**Unknown**

$$m_{\text{NH}_3} = ? \text{ g}$$

### 2 SOLVE FOR THE UNKNOWN

Determine how many liters of gaseous ammonia will be made from 5.00 L of nitrogen gas.

$$\frac{1 \text{ volume N}_2}{2 \text{ volumes NH}_3}$$

$$5.00 \text{ L N}_2 \left( \frac{2 \text{ volumes NH}_3}{1 \text{ volume N}_2} \right) = 10.0 \text{ L NH}_3$$

Find the volume ratio for  $\text{N}_2$  and  $\text{NH}_3$  using the balanced equation.

Multiply the known volume of  $\text{N}_2$  by the volume ratio to find the volume of  $\text{NH}_3$ .

Use the ideal gas law. Solve for  $n$ , and calculate the number of moles of  $\text{NH}_3$ .

$$PV = nRT$$

$$n = \frac{PV}{RT}$$

$$n = \frac{(3.00 \text{ atm})(10.0 \text{ L})}{(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(298 \text{ K})}$$

$$n = \frac{(3.00 \text{ atm})(10.0 \text{ L})}{(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(298 \text{ K})} = 1.23 \text{ mol NH}_3$$

State the ideal gas law.

Solve for  $n$ .

Substitute  $P = 3.00 \text{ atm}$ ,  $V_{\text{NH}_3} = 10.0 \text{ L}$ , and  $T = 298 \text{ K}$ .

Multiply and divide numbers and units.

$$M = \left( \frac{1 \text{ N-atom} \times 14.01 \text{ amu}}{1 \text{ N-atom}} \right) + \left( \frac{3 \text{ H-atoms} \times 1.01 \text{ amu}}{1 \text{ H-atom}} \right)$$
$$= 17.04 \text{ amu}$$

Find the molecular mass of  $\text{NH}_3$ .

$$M = 17.04 \text{ g/mol}$$

Express molar mass in units of g/mol.

Convert moles of ammonia to grams of ammonia.

$$1.23 \text{ mol NH}_3 \times \frac{17.04 \text{ g NH}_3}{1 \text{ mol NH}_3} = 21.0 \text{ g NH}_3$$

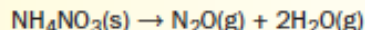
Use the molar mass of ammonia as a conversion factor.

### 3 EVALUATE THE ANSWER

To check your answer, calculate the volume of reactant nitrogen at STP. Then, use molar volume and the mole ratio between  $\text{N}_2$  and  $\text{NH}_3$  to determine how many moles of  $\text{NH}_3$  were produced. The unit of the answer is grams, a unit of mass. There are three significant figures.

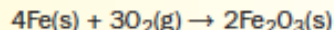
## APPLICATIONS

42. Ammonium nitrate is a common ingredient in chemical fertilizers. Use the reaction shown to calculate the mass of solid ammonium nitrate that must be used to obtain 0.100 L of dinitrogen monoxide gas at STP.



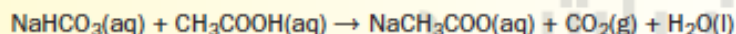
43. When solid calcium carbonate ( $\text{CaCO}_3$ ) is heated, it decomposes to form solid calcium oxide ( $\text{CaO}$ ) and carbon dioxide gas ( $\text{CO}_2$ ). How many liters of carbon dioxide will be produced at STP if 2.38 kg of calcium carbonate reacts completely?

44. When iron rusts, it undergoes a reaction with oxygen to form iron(III) oxide.



Calculate the volume of oxygen gas at STP that is required to completely react with 52.0 g of iron.

45. **Challenge** An excess of acetic acid is added to 28 g of sodium bicarbonate at  $25^\circ\text{C}$  and 1 atm pressure. During the reaction, the gas cools to  $20^\circ\text{C}$ . What volume of carbon dioxide will be produced? The balanced equation for the reaction is shown below.



42. 0.357 g  $\text{NH}_4\text{NO}_3$

43. 533 L  $\text{CO}_2$

44. 15.6 L  $\text{O}_2$

45. 7.9 L  $\text{CO}_2$

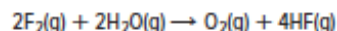
## Section 3 Review

### SECTION 3 REVIEW

#### Section Summary

- The coefficients in a balanced chemical equation specify volume ratios for gaseous reactants and products.
- The gas laws can be used along with balanced chemical equations to calculate the amount of a gaseous reactant or product in a reaction.

46. **MAIN IDEA Explain** When fluorine gas combines with water vapor, the following reaction occurs.



If the reaction starts with 2 L of fluorine gas, how many liters of water vapor react with the fluorine, and how many liters of oxygen and hydrogen fluoride are produced?

47. **Analyze** Is the volume of a gas directly or inversely proportional to the number of moles of a gas at constant temperature and pressure? Explain.

48. **Calculate** One mole of a gas occupies a volume of 22.4 L at STP. Calculate the temperature and pressure conditions needed to fit 2 mol of a gas into a volume of 22.4 L.

49. **Interpret Data** Ethene gas ( $\text{C}_2\text{H}_4$ ) reacts with oxygen to form carbon dioxide and water. Write a balanced equation for this reaction, then find the mole ratios of substances on each side of the equation.

### SECTION 3 REVIEW

46. 2 L  $\text{H}_2\text{O}$ , 1 L  $\text{O}_2$ , and 4 L  $\text{HF}$

47. Directly proportional; as the amount of gas increases, so does volume.

48. Student answers will vary. Temperature can be halved or pressure doubled or a combination of lowering temperature and increasing pressure.

49.  $\text{C}_2\text{H}_4(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2 + 2\text{H}_2\text{O}$

The mole ratio of ethene to oxygen is 1:3; the mole ratio of carbon dioxide to water is 2:2.

## SECTION 1 The Gas Laws

**MAINIDEA** For a fixed amount of gas, a change in one variable—pressure, temperature, or volume—affects the other two.

- Boyle's law states that the volume of a fixed amount of gas is inversely proportional to its pressure at constant temperature.

$$P_1V_1 = P_2V_2$$

- Charles's law states that the volume of a fixed amount of gas is directly proportional to its Kelvin temperature at constant pressure.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

- Gay-Lussac's law states that the pressure of a fixed amount of gas is directly proportional to its Kelvin temperature at constant volume.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

- The combined gas law relates pressure, temperature, and volume in a single statement.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

### VOCABULARY

- Boyle's law
- absolute zero
- Charles's law
- Gay-Lussac's law
- combined gas law

## SECTION 2 The Ideal Gas Law

**MAINIDEA** The ideal gas law relates the number of particles to pressure, temperature, and volume.

- Avogadro's principle states that equal volumes of gases at the same pressure and temperature contain equal numbers of particles.
- The ideal gas law relates the amount of a gas present to its pressure, temperature, and volume.

$$PV = nRT$$

- The ideal gas law can be used to find molar mass if the mass of the gas is known or the density of the gas if its molar mass is known.

$$M = \frac{mRT}{PV} \quad D = \frac{MP}{RT}$$

- At very high pressures and very low temperatures, real gases behave differently than ideal gases.

### VOCABULARY

- Avogadro's principle
- molar volume
- standard temperature and pressure (STP)
- ideal gas constant (R)
- ideal gas law

## SECTION 3 Gas Stoichiometry

**MAINIDEA** When gases react, the coefficients in the balanced chemical equation represent both molar amounts and relative volumes.

- The coefficients in a balanced chemical equation specify volume ratios for gaseous reactants and products.
- The gas laws can be used along with balanced chemical equations to calculate the amount of a gaseous reactant or product in a reaction.



## Chapter 9 Review and MS

### SECTION 1

#### Mastering Concepts

50. State Boyle's law, Charles's law, Gay-Lussac's law, and the combined gas law in words and equations.
51. If two variables are inversely proportional, what happens to the value of one as the value of the other increases?
52. If two variables are directly proportional, what happens to the value of one as the value of the other increases?
53. List the standard conditions for gas measurements.
54. Identify the units most commonly used for  $P$ ,  $V$ , and  $T$ .

#### Mastering Problems

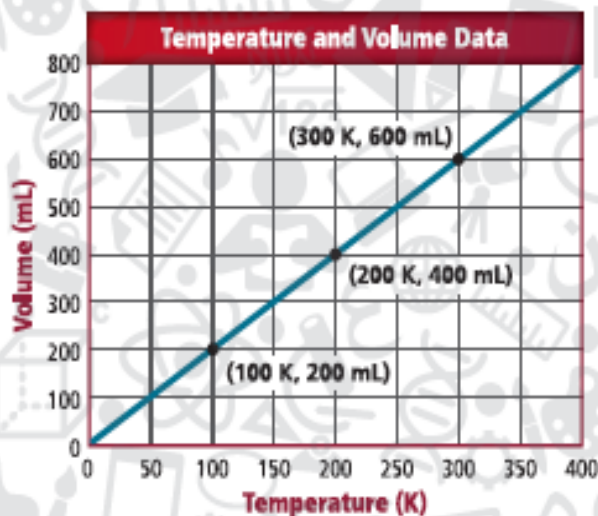


Figure 13

55. Use Charles's law to determine the accuracy of the data plotted in Figure 13.
56. **Weather Balloons** A weather balloon is filled with helium that occupies a volume of  $5.00 \times 10^4$  L at 0.995 atm and  $32.0^\circ\text{C}$ . After it is released, it rises to a location where the pressure is 0.720 atm and the temperature is  $-12.0^\circ\text{C}$ . What is the volume of the balloon at the new location?

**50.** Boyle's law: the volume of a given amount of gas held at a constant temperature varies inversely with pressure,  $P_1V_1 = P_2V_2$ ; Charles's law: the volume of a given mass of gas is directly proportional to its kelvin temperature at constant pressure,  $V_1/T_1 = V_2/T_2$ ; Gay-Lussac's law: the pressure of a given mass of gas varies directly with the kelvin temperature when the volume remains constant,  $P_1/T_1 = P_2/T_2$ ; combined gas law: states the relationships between pressure, volume, and temperature of a fixed amount of gas,  $P_1V_1/T_1 = P_2V_2/T_2$

51. One variable always decreases as the other increases.
52. One variable always increases as the other increases.
53.  $T = 0.00^\circ\text{C}$  (273 K) and  $P = 1.00$  atm
54. atm for pressure, kelvins for temperature, and L for volume
55. Charles's law states that the volume of a given mass of gas is directly proportional to temperature. The graphed data follow this law because doubling the temperature doubles the volume. Therefore, the data are accurate.

**56.**  $5.91 \times 10^4$  L



57. Use Boyle's, Charles's, or Gay-Lussac's law to calculate the missing value in each of the following.

- a.  $V_1 = 2.0 \text{ L}$ ,  $P_1 = 0.82 \text{ atm}$ ,  $V_2 = 1.0 \text{ L}$ ,  $P_2 = ?$
- b.  $V_1 = 250 \text{ mL}$ ,  $T_1 = ?$ ,  $V_2 = 400 \text{ mL}$ ,  $T_2 = 298 \text{ K}$
- c.  $V_1 = 0.55 \text{ L}$ ,  $P_1 = 740 \text{ mm Hg}$ ,  $V_2 = 0.80 \text{ L}$ ,  $P_2 = ?$

58. **Hot-Air Balloons** A sample of air occupies  $2.50 \text{ L}$  at a temperature of  $22.0^\circ\text{C}$ . What volume will this sample occupy inside a hot-air balloon at a temperature of  $43.0^\circ\text{C}$ ? Assume that the pressure inside the balloon remains constant.

59. What is the pressure of a fixed volume of hydrogen gas at  $30.0^\circ\text{C}$  if it has a pressure of  $1.11 \text{ atm}$  at  $15.0^\circ\text{C}$ ?

- 57. a.  $1.6 \text{ atm}$
- b.  $200 \text{ K}$
- c.  $510 \text{ mm Hg}$

58.  $2.68 \text{ L}$

59.  $1.17 \text{ atm}$

60.  $74.8 \text{ kPa}$

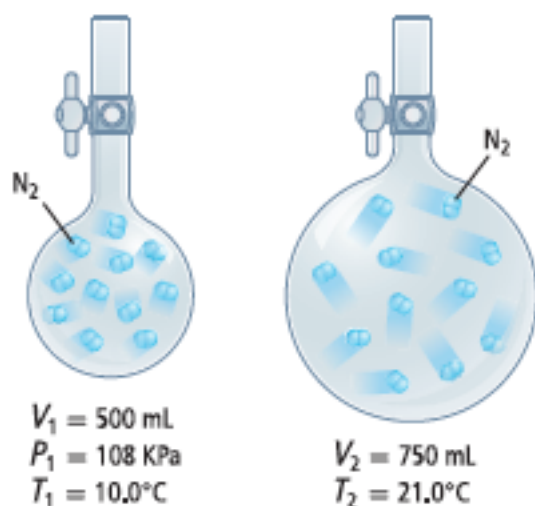


Figure 14

60. A sample of nitrogen gas is transferred to a larger flask, as shown in Figure 14. What is the pressure of nitrogen in the second flask?

## SECTION 2

- 61. State Avogadro's principle.
- 62. State the ideal gas law.
- 63. What volume is occupied by 1 mol of a gas at STP? What volume does 2 mol occupy at STP?
- 64. Define the term *ideal gas*, and explain why there are no true ideal gases in nature.
- 65. List two conditions under which a gas is least likely to behave ideally.
- 66. What units must be used to express the temperature in the equation for the ideal gas law? Explain.

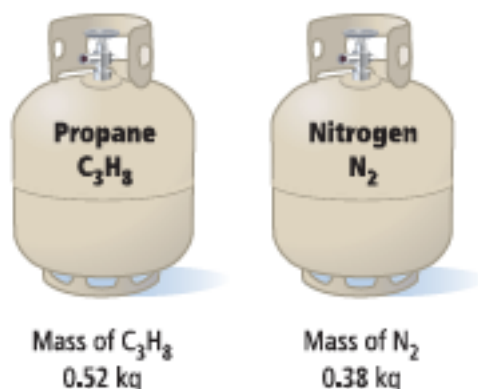
### Mastering Problems

- 67. **Home Fuel** Propane ( $\text{C}_3\text{H}_8$ ) is a gas commonly used as a home fuel for cooking and heating.
  - a. Calculate the volume that 0.540 mol of propane occupies at STP.
  - b. Think about the size of this volume and the amount of propane that it contains. Why do you think propane is usually liquefied before it is transported?
- 68. **Careers in Chemistry** A physical chemist measured the lowest pressure achieved in a laboratory—about  $1.0 \times 10^{-15}$  mm Hg. How many molecules of gas are present in a 1.00 L sample at that pressure if the sample's temperature is  $22.0^\circ\text{C}$ ?
- 69. Calculate the number of moles of  $\text{O}_2$  gas held in a sealed, 2.00 L tank at 3.50 atm and  $25.0^\circ\text{C}$ . How many moles would be in the tank if the temperature was raised to  $49.0^\circ\text{C}$  and the pressure remained constant?

- 61. At a fixed temperature and pressure, equal volumes of any ideal gas contain equal numbers of particles.
- 62. The ideal gas law describes the physical behavior of an ideal gas in terms of the pressure, volume, temperature, and number of moles of gas present:  $PV=nRT$ .
- 63. 22.4 L; 44.8 L

- 64. An ideal gas is one whose particles take up no space and have no intermolecular attractive forces and it follows the gas laws under all conditions of temperature and pressure. No gas is truly ideal because all gas particles have some volume and are subject to intermolecular interactions.
- 65. high pressure and low temperature
- 66. kelvins;  $V$  is not directly proportional to Celsius temperature.
- 67. a. 12.1 L  
b. Propane occupies a much smaller volume when liquefied.
- 68.  $3.3 \times 10^4$  molecules
- 69. 0.286 mol; 0.265 mol

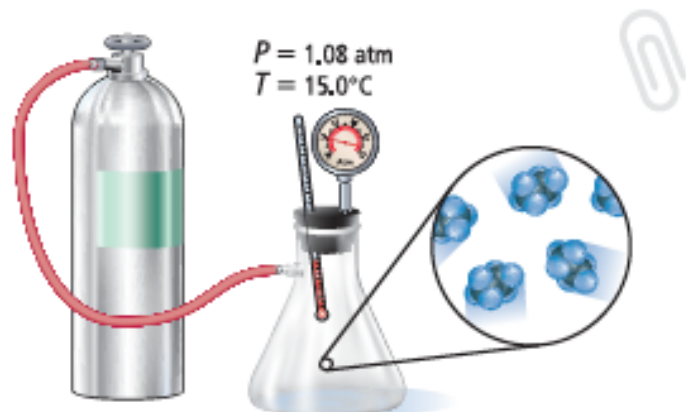
- 70. Perfumes** Geraniol is a compound found in rose oil that is used in perfumes. What is the molar mass of geraniol if its vapor has a density of 0.480 g/L at a temperature of 260.0°C and a pressure of 0.140 atm?
- 71.** Find the volume that 42 g of carbon monoxide gas occupies at STP.
- 72.** Determine the density of chlorine gas at 22.0°C and 1.00 atm.



■ **Figure 15**

- 73.** Which of the gases in **Figure 15** occupies the greatest volume at STP? Explain your answer.
- 74.** If the containers in **Figure 15** each hold 4.00 L, what is the pressure inside each? Assume ideal behavior.

- 70.**  $1.50 \times 10^2$  g/mol
- 71.** 34 L
- 72.** 2.93 g/L
- 73.** The  $N_2$  occupies the greatest volume at STP. The  $N_2$  occupies 310 L, while the  $C_3H_8$  occupies only 260 L.
- 74.** propane: 66.1 atm; nitrogen: 76.2 atm



■ Figure 16

75. A 2.00 L flask is filled with ethane gas ( $\text{C}_2\text{H}_6$ ) from a small cylinder, as shown in Figure 16. What is the mass of the ethane in the flask?
76. What is the density of a sample of nitrogen gas ( $\text{N}_2$ ) that exerts a pressure of 5.30 atm in a 3.50 L container at  $125^\circ\text{C}$ ?
77. How many moles of helium gas ( $\text{He}$ ) would be required to fill a 22 L container at a temperature of  $35^\circ\text{C}$  and a pressure of 3.1 atm?
78. Before a reaction, two gases share a container at a temperature of 200 K. After the reaction, the product is in the same container at a temperature of 400 K. If both  $V$  and  $P$  are constant, what must be true of  $n$ ?

75. 2.75 g

76. 4.55 g/L

77. 2.7 mol

78. With a constant volume and pressure and doubling of the temperature, the number of moles would be half the initial number of moles

## SECTION 3

### Mastering Concepts

79. Why must an equation be balanced before using it to determine the volumes of gases involved in a reaction?
80. It is not necessary to consider temperature and pressure when using a balanced equation to determine relative gas volume. Why?
81. What information do you need to solve a volume-mass problem that involves gases?
82. Explain why the coefficients in a balanced chemical equation represent not only molar amounts but also relative volumes for gases.
83. Do the coefficients in a balanced chemical equation represent volume ratios for solids and liquids? Explain.

### Mastering Problems

**84. Ammonia Production** Ammonia is often formed by reacting nitrogen and hydrogen gases. How many liters of ammonia gas can be formed from 13.7 L of hydrogen gas at 93.0°C and a pressure of 40.0 kPa?

**85.** A 6.5 L sample of hydrogen sulfide is treated with a catalyst to promote the reaction shown below.



If the  $\text{H}_2\text{S}$  reacts completely at 2.0 atm and 290 K, how much water vapor, in grams, is produced?

**86.** To produce 15.4 L of nitrogen dioxide at 310 K and 2.0 atm, how many liters of nitrogen gas and oxygen gas are required?

**87.** Use the reaction shown below to answer these questions.



- a. What is the volume ratio of carbon monoxide to carbon dioxide in the balanced equation?
- b. If 42.7 g of CO is reacted completely at STP, what volume of  $\text{N}_2$  gas will be produced?

**79.** Equation coefficients represent ratios between gas volumes in the reaction.

**80.** Temperature and pressure are the same for each gas involved in the reaction. These conditions affect each gas in the same way.

**81.** balanced equation, at least one mass or volume value for a reactant or product, and  $P$  and  $T$  conditions under which gas volumes have been measured

**82.** Avogadro's principle states that equal volumes of gases at the same temperature and pressure contain equal numbers of particles (or an equal number of moles). Therefore, the coefficients also represent the relative volumes of the gases.

**83.** No, this relationship only applies to gases that behave like ideal gases.

**84.** 9.13 L

**85.** 9.9 g

**86.**  $\text{N}_2$  gas: 7.7 L;  $\text{O}_2$  gas: 15.4 L

**87. a.** 1:1

**b.** 17.1 L



- 88.** When 3.00 L of propane gas is completely combusted to form water vapor and carbon dioxide at 350°C and 0.990 atm, what mass of water vapor results?
- 89.** When heated, solid potassium chlorate ( $\text{KClO}_3$ ) decomposes to form solid potassium chloride and oxygen gas. If 20.8 g of potassium chlorate decomposes, how many liters of oxygen gas will form at STP?
- 90. Welding** The gas acetylene, often used for welding, burns according to the following equation.



If you have a 10.0 L tank of acetylene at 25.0°C and 1.00 atm pressure, how many moles of  $\text{CO}_2$  will be produced if you burn all the acetylene in the tank?

**88.** 4.2 g

**89.** 5.71 L

**90.** 0.817 mol

# Chapter 10 “Mixtures and Solutions”

## Section 1 Types of Mixtures

### Students Learning outcomes:

- Differentiate qualitatively and quantitatively between the different types of mixtures and solutions.
- Compare the formation process of solutions that are produced by dissolving ionic and molecular compounds in water, and of solutions that are produced by dissolving non-polar solutes in non-polar solvents.

### New vocabulary-

- **Solute:** a substance dissolved in a solution.
- **Suspension:** is a mixture containing particles that settle out if left undisturbed.



- **Colloid:** A heterogeneous mixture of intermediate sized particles (between atomic-scale size of solution particles and the size of suspension particles).
- **Brownian motion:** The dispersed particles of liquid colloids make jerky, random movements. This erratic movement of colloid particles.
- **Tyndall effect:** A phenomena where dispersed colloid particles scatter light.
- **Soluble:** A substance that dissolves in a solvent.
- **Miscible:** Two liquids that are soluble in each other in any proportion.
- **Insoluble:** A substance that does not dissolve in a solvent.
- **Immiscible:** liquids that can be mixed together but separate shortly after.

**Table 1 Types of Colloids**

Category	Example	Dispersed Particles	Dispersing Medium
Solid sol	colored gems	solid	solid
Sol	blood, gelatin	solid	liquid
Solid emulsion	butter, cheese	liquid	solid
Emulsion	milk, mayonnaise	liquid	liquid
Solid foam	marshmallow, soaps that float	gas	solid
Foam	whipped cream, beaten egg white	gas	liquid
Solid aerosol	smoke, dust in air	solid	gas
Liquid aerosol	spray deodorant, fog, clouds	liquid	gas

## Section 1 Review

### SECTION 1 REVIEW

#### Section Summary

- The individual substances in a heterogeneous mixture remain distinct.
- Two types of heterogeneous mixtures are suspensions and colloids.
- Brownian motion is the erratic movement of colloid particles.
- Colloids exhibit the Tyndall effect.
- A solution can exist as a gas, a liquid, or a solid, depending on the solvent.
- Solutes in a solution can be gases, liquids, or solids.

- 1. MAIN IDEA Explain** Use the properties of seawater to describe the characteristics of mixtures.
- 2. Distinguish** between suspensions and colloids.
- 3. Identify** the various types of solutions. Describe the characteristics of each type of solution.
- 4. Explain** Use the Tyndall effect to explain why it is more difficult to drive through fog using high beams than using low beams.
- 5. Describe** different types of colloids.
- 6. Explain** Why do dispersed colloid particles stay dispersed?
- 7. Summarize** What causes Brownian motion?
- 8. Compare and Contrast** Make a table that compares the properties of suspensions, colloids, and solutions.

### SECTION 1 REVIEW

- Answers will vary but might include that seawater is a heterogeneous mixture with dirt and mud particles, and it is a homogeneous mixture with dissolved substances.
- Suspension particles are larger than colloidal particles. Suspension particles settle out of the mixture, whereas colloidal particles do not.
- All solutions are homogeneous mixtures containing two or more substances. Solutions might be liquid, solid, or gas. Solution types are identified in Table 2.
- High beams are aimed farther down the road than low beams. Because the fog scatters light, there is less light from the high beams to illuminate the road than from the low beams. Also,

- because the high beams are aimed more directly into the fog, more of their light is reflected back toward the driver, making it more difficult to see.
- Refer to Table 1 for descriptions of colloid types.
- The particles do not settle out because they have polar or charged layers surrounding them. These layers repel each other and prevent the particles from settling or separating.
- Collisions of particles of the dispersion medium with the dispersed particles result in Brownian motion.
- Student tables will vary, but should include particle size, if the particles settle out, and if the particles display the Tyndall effect.

## Section 2 Solution Concentration

### Students Learning outcomes:

- Solve problems related to the concentration of solutions by performing calculation involving moles and express the results in various units (e.g.mol/l (molarity), mol/kg (molality), part per million (ppm), percent by mass, percent by volume).
- Prepare practically solutions with specific concentrations by dissolving a solid solute in a solvent or by diluting the concentrated solution.

### New vocabulary-

- **Solvent:** the substance that dissolves a solute to form a solution
- **Concentration: (of a solution)** is a measure of how much solute is dissolved in a specific amount of solvent or solution.

Table 3 Concentration Ratios	
Concentration Description	Ratio
Percent by mass	$\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$
Percent by volume	$\frac{\text{volume of solute}}{\text{volume of solution}} \times 100$
Molarity	$\frac{\text{moles of solute}}{\text{liter of solution}}$
Molality	$\frac{\text{moles of solute}}{\text{kilogram of solvent}}$
Mole fraction	$\frac{\text{moles of solute}}{\text{moles of solute} + \text{moles of solvent}}$

- **Molarity:** is the number of moles of solute dissolved per liter of solution.

### Molarity

$$\text{molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

The molarity of a solution equals the moles of solute divided by the liters of solution.

- **Molality:** the ratio of the number of moles of solute dissolved in 1 kg of solvent. The unit m is read as molal.

### Molality

$$\text{molality } (m) = \frac{\text{moles of solute}}{\text{kg of solvent}}$$

The molality of a solution equals the moles of solute divided by kilograms of solvent.

- **Mole fraction:** the ratio of the number of moles of solute or solvent in solution to the total number of moles of solute and solvent.

### Mole Fraction

$$X_A = \frac{n_A}{n_A + n_B} \quad X_B = \frac{n_B}{n_A + n_B}$$

$X_A$  and  $X_B$  represent the mole fractions of each substance.

$n_A$  and  $n_B$  represent the number of moles of each substance.

A mole fraction equals the number of moles of solute or solvent in a solution divided by the total number of moles of solute and solvent.



## Worked Examples -

**Percent by mass** The percent by mass is the ratio of the solute's mass to the solution's mass expressed as a percent. The mass of the solution equals the sum of the masses of the solute and the solvent.

### Percent by Mass

$$\text{percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

Percent by mass equals the mass of the solute divided by the mass of the whole solution, multiplied by 100.

### EXAMPLE 1

**CALCULATE PERCENT BY MASS** In order to maintain a sodium chloride (NaCl) concentration similar to ocean water, an aquarium must contain 3.6 g NaCl per 100.0 g of water. What is the percent by mass of NaCl in the solution?

#### 1 ANALYZE THE PROBLEM

You are given the amount of sodium chloride dissolved in 100.0 g of water. The percent by mass of a solute is the ratio of the solute's mass to the solution's mass, which is the sum of the masses of the solute and the solvent.

##### Known

mass of solute = 3.6 g NaCl

mass of solvent = 100.0 g H<sub>2</sub>O

##### Unknown

percent by mass = ?

#### 2 SOLVE FOR THE UNKNOWN

Find the mass of the solution.

mass of solution = grams of solute + grams of solvent

mass of solution = 3.6 g + 100.0 g = 103.6 g *Substitute mass of solute = 3.6 g, and mass of solvent = 100.0 g.*

Calculate the percent by mass.

$$\text{percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

*State the equation for percent by mass.*

$$\text{percent by mass} = \frac{3.6 \text{ g}}{103.6 \text{ g}} \times 100 = 3.5\%$$

*Substitute mass of solute = 3.6 g, and mass of solution = 103.6 g.*

#### 3 EVALUATE THE ANSWER

Because only a small mass of sodium chloride is dissolved per 100.0 g of water, the percent by mass should be a small value, which it is. The mass of sodium chloride was given with two significant figures; therefore, the answer is also expressed with two significant figures.

### APPLICATIONS

- What is the percent by mass of NaHCO<sub>3</sub> in a solution containing 20.0 g of NaHCO<sub>3</sub> dissolved in 600.0 mL of H<sub>2</sub>O?
- You have 1500.0 g of a bleach solution. The percent by mass of the solute sodium hypochlorite (NaOCl) is 3.62%. How many grams of NaOCl are in the solution?
- In Question 10, how many grams of solvent are in the solution?
- Challenge** The percent by mass of calcium chloride in a solution is found to be 2.65%. If 50.0 g of calcium chloride is used, what is the mass of the solution?

- 3.23%
- 54.3 g
- 1445.7 g
- $1.89 \times 10^3 \text{ g}$

### Percent by Volume

$$\text{percent by volume} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$$

Percent by volume equals the volume of solute divided by the volume of the solution, multiplied by 100.

### APPLICATIONS

13. What is the percent by volume of ethanol in a solution that contains 35 mL of ethanol dissolved in 155 mL of water?
14. What is the percent by volume of isopropyl alcohol in a solution that contains 24 mL of isopropyl alcohol in 1.1 L of water?
15. **Challenge** If 18 mL of methanol is used to make an aqueous solution that is 15% methanol by volume, how many milliliters of solution is produced?

13. 18%

14. 2.1%

15. 120 mL

## EXAMPLE 2

**CALCULATING MOLARITY** A 100.5 mL intravenous (IV) solution contains 5.10 g of glucose ( $\text{C}_6\text{H}_{12}\text{O}_6$ ). What is the molarity of this solution? The molar mass of glucose is 180.16 g/mol.

### 1 ANALYZE THE PROBLEM

You are given the mass of glucose dissolved in a volume of water. The molarity of the solution is the ratio of moles of solute per liter of solution.

#### Known

mass of solute = 5.10 g  $\text{C}_6\text{H}_{12}\text{O}_6$

molar mass of  $\text{C}_6\text{H}_{12}\text{O}_6$  = 180.16 g/mol

volume of solution = 100.5 mL

#### Unknown

solution concentration = ? M

### 2 SOLVE FOR THE UNKNOWN

Calculate the number of moles of  $\text{C}_6\text{H}_{12}\text{O}_6$ .

$$(5.10 \text{ g } \text{C}_6\text{H}_{12}\text{O}_6) \left( \frac{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{180.16 \text{ g } \text{C}_6\text{H}_{12}\text{O}_6} \right)$$

$$= 0.0283 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6$$

Multiply grams of  $\text{C}_6\text{H}_{12}\text{O}_6$  by the molar mass of  $\text{C}_6\text{H}_{12}\text{O}_6$ .

Convert the volume of  $\text{H}_2\text{O}$  to liters.

$$(100.5 \text{ mL solution}) \left( \frac{1 \text{ L}}{1000 \text{ mL}} \right) = 0.1005 \text{ L solution}$$

Use the conversion factor 1 L/1000 mL.

Solve for the molarity.

$$M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

State the molarity equation.

$$M = \left( \frac{0.0283 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{0.1005 \text{ L solution}} \right)$$

Substitute moles of  $\text{C}_6\text{H}_{12}\text{O}_6$  = 0.0283 and volume of solution = liters of solution = 0.1005 L.

$$M = \left( \frac{0.0282 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{1 \text{ L solution}} \right) = 0.282M$$

Divide numbers and units.

### 3 EVALUATE THE ANSWER

The molarity value will be small because only a small mass of glucose was dissolved in the solution. The mass of glucose used in the problem has three significant figures; therefore, the value of the molarity also has three significant figures.

### APPLICATIONS

16. What is the molarity of an aqueous solution containing 40.0 g of glucose ( $C_6H_{12}O_6$ ) in 1.5 L of solution?
17. Calculate the molarity of 1.60 L of a solution containing 1.55 g of dissolved KBr.
18. What is the molarity of a bleach solution containing 9.5 g of NaOCl per liter of bleach?
19. **Challenge** How much calcium hydroxide ( $Ca(OH)_2$ ), in grams, is needed to produce 1.5 L of a 0.25M solution?

- |                            |           |
|----------------------------|-----------|
| 16. 0.15M                  | 18. 0.13M |
| 17. $8.13 \times 10^{-3}M$ | 19. 28 g  |

### APPLICATIONS

20. How many grams of  $CaCl_2$  would be dissolved in 1.0 L of a 0.10M solution of  $CaCl_2$ ?
21. How many grams of  $CaCl_2$  should be dissolved in 500.0 mL of water to make a 0.20M solution of  $CaCl_2$ ?
22. What mass of NaOH is in 250 mL of a 3.0M NaOH solution?
23. **Challenge** What volume of ethanol ( $C_2H_5OH$ ) is in 100.0 mL of 0.15M solution? The density of ethanol is 0.7893 g/mL.

- |                         |
|-------------------------|
| 20. 11 g                |
| 21. 11 g                |
| 22. $3.0 \times 10^1 g$ |
| 23. 0.87 mL             |

## Dilution Equation

$$M_1V_1 = M_2V_2$$

*M* represents molarity.  
*V* represents volume.

For a given amount of solute, the product of the molarity and volume of the stock solution equals the product of the molarity and the volume of the dilute solution.

### EXAMPLE 3

**DILUTING STOCK SOLUTIONS** If you know the concentration and volume of the solution you want to prepare, you can calculate the volume of stock solution you will need. What volume, in milliliters, of 2.00M calcium chloride ( $\text{CaCl}_2$ ) stock solution would you use to make 0.50 L of 0.300M calcium chloride solution?

#### 1 ANALYZE THE PROBLEM

You are given the molarity of a stock solution of  $\text{CaCl}_2$  and the volume and molarity of a dilute solution of  $\text{CaCl}_2$ . Use the relationship between molarities and volumes to find the volume, in liters, of the stock solution required. Then, convert the volume to milliliters.

##### Known

$$\begin{aligned}M_1 &= 2.00M \text{ CaCl}_2 \\M_2 &= 0.300M \\V_2 &= 0.50 \text{ L}\end{aligned}$$

##### Unknown

$$V_1 = ? \text{ mL } 2.00M \text{ CaCl}_2$$

#### 2 SOLVE FOR THE UNKNOWN

Solve the molarity-volume relationship for the volume of the stock solution  $V_1$ .

$$M_1V_1 = M_2V_2$$

State the dilution equation.

$$V_1 = V_2 \left( \frac{M_2}{M_1} \right)$$

Solve for  $V_1$ .

$$V_1 = (0.50 \text{ L}) \left( \frac{0.300M}{2.00M} \right)$$

Substitute  $M_1 = 2.00M$ ,  
 $M_2 = 0.300M$ , and  $V_2 = 0.50 \text{ L}$ .

$$V_1 = (0.50 \text{ L}) \left( \frac{0.300M}{2.00M} \right) = 0.075 \text{ L}$$

Multiply and divide numbers  
and units.

$$V_1 = (0.075 \text{ L}) \left( \frac{1000 \text{ mL}}{1 \text{ L}} \right) = 75 \text{ mL}$$

Convert to milliliters using the  
conversion factor 1000 mL/1 L.

To make the dilution, measure out 75 mL of the stock solution and dilute it with enough water to make the final volume 0.50 L.

#### 3 EVALUATE THE ANSWER

The volume  $V_1$  was calculated, and then its value was converted to milliliters. This volume should be less than the final volume of the dilute solution, and it is. Of the given information,  $V_2$  had the fewest number of significant figures, with two. Thus, the volume  $V_1$  should also have two significant figures, and it does.



## APPLICATIONS

24. What volume of a 3.00M KI stock solution would you use to make 0.300 L of a 1.25M KI solution?
25. How many milliliters of a 5.0M  $\text{H}_2\text{SO}_4$  stock solution would you need to prepare 100.0 mL of 0.25M  $\text{H}_2\text{SO}_4$ ?
26. **Challenge** If 0.50 L of 5.00M stock solution of HCl is diluted to make 2.0 L of solution, how much HCl, in grams, is in the solution?

24. 125 mL  
25. 5.0 mL  
26. 91 g

## EXAMPLE 4

**CALCULATING MOLALITY** In the lab, a student adds 4.5 g of sodium chloride (NaCl) to 100.0 g of water. Calculate the molality of the solution.

### 1 ANALYZE THE PROBLEM

You are given the mass of solute and solvent. Determine the number of moles of solute. Then, you can calculate the molality.

#### Known

mass of water ( $\text{H}_2\text{O}$ ) = 100.0 g

mass of sodium chloride (NaCl) = 4.5 g

#### Unknown

$m = ? \text{ mol/kg}$

### 2 SOLVE FOR THE UNKNOWN

$$4.5 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 0.077 \text{ mol NaCl}$$

Calculate the number of moles of solute.

$$100.0 \text{ g H}_2\text{O} \times \frac{1 \text{ kg H}_2\text{O}}{1000 \text{ g H}_2\text{O}} = 0.1000 \text{ kg H}_2\text{O}$$

Convert the mass of  $\text{H}_2\text{O}$  from grams to kilograms using the factor 1 kg/1000 g.

Substitute the known values into the expression for molality, and solve.

$$m = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$$

Write the equation for molality.

$$m = \frac{0.077 \text{ mol NaCl}}{0.1000 \text{ kg H}_2\text{O}} = 0.77 \text{ mol/kg}$$

Substitute moles of solute = 0.077 mol NaCl, kilograms of solvent = 0.1000 kg  $\text{H}_2\text{O}$ .

### 3 EVALUATE THE ANSWER

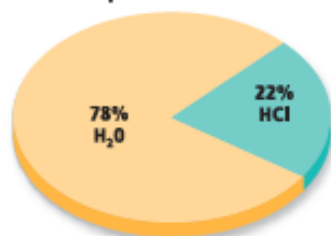
Because there was less than one-tenth of a mole of solute present in one-tenth of a kilogram of water, the molality should be less than one, and it is. The mass of sodium chloride was given with two significant figures; therefore, the molality is also expressed with two significant figures.

## APPLICATIONS

27. What is the molality of a solution containing 10.0 g of  $\text{Na}_2\text{SO}_4$  dissolved in 1000.0 g of water?
28. **Challenge** How much  $(\text{Ba}(\text{OH})_2)$ , in grams, is needed to make a 1.00m aqueous solution?

27. 0.0704m  
28. 171 g per kg of water

### Hydrochloric Acid in Aqueous Solution



$$X_{\text{HCl}} + X_{\text{H}_2\text{O}} = 1.00$$

$$0.22 + 0.78 = 1.00$$

**Figure 8** The mole fraction expresses the number of moles of solute and solvent relative to the total number of moles of solution. Each mole fraction can be thought of as a percent. For example, the mole fraction of water ( $X_{\text{H}_2\text{O}}$ ) is 0.78, which is equivalent to saying the solution contains 78% water (on a mole basis).

**Mole fraction** If you know the number of moles of solute and solvent, you can also express the concentration of a solution as a **mole fraction**—the ratio of the number of moles of solute or solvent in solution to the total number of moles of solute and solvent, as shown in **Figure 8**.

The symbol  $X$  is commonly used for mole fraction, with a subscript to indicate the solvent or solute. The mole fraction for the solvent ( $X_A$ ) and the mole fraction for the solute ( $X_B$ ) can be expressed as follows.

#### Mole Fraction

$$X_A = \frac{n_A}{n_A + n_B} \quad X_B = \frac{n_B}{n_A + n_B}$$

$X_A$  and  $X_B$  represent the mole fractions of each substance.  
 $n_A$  and  $n_B$  represent the number of moles of each substance.

A mole fraction equals the number of moles of solute or solvent in a solution divided by the total number of moles of solute and solvent.

For example, suppose a hydrochloric acid solution contains 36 g of HCl and 64 g of  $\text{H}_2\text{O}$ . To convert these masses to moles, you would use the molar masses as conversion factors.

$$n_{\text{HCl}} = 36 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.5 \text{ g HCl}} = 0.99 \text{ mol HCl}$$

$$n_{\text{H}_2\text{O}} = 64 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} = 3.6 \text{ mol H}_2\text{O}$$

The mole fractions of HCl and water can be expressed as follows.

$$X_{\text{HCl}} = \frac{n_{\text{HCl}}}{n_{\text{HCl}} + n_{\text{H}_2\text{O}}} = \frac{0.99 \text{ mol HCl}}{0.99 \text{ mol HCl} + 3.6 \text{ mol H}_2\text{O}} = 0.22$$

$$X_{\text{H}_2\text{O}} = \frac{n_{\text{H}_2\text{O}}}{n_{\text{HCl}} + n_{\text{H}_2\text{O}}} = \frac{3.6 \text{ mol H}_2\text{O}}{0.99 \text{ mol HCl} + 3.6 \text{ mol H}_2\text{O}} = 0.78$$

### APPLICATIONS

29. What is the mole fraction of NaOH in an aqueous solution that contains 22.8% NaOH by mass?
30. **Challenge** If the mole fraction of sulfuric acid ( $\text{H}_2\text{SO}_4$ ) in an aqueous solution is 0.325, what is the percent by mass of  $\text{H}_2\text{SO}_4$ ?

29. 0.118

30. 72.3%

## Section 2 Solution Concentration

### SECTION 2 REVIEW

#### Section Summary

- Concentrations can be measured qualitatively and quantitatively.
- Molarity is the number of moles of solute dissolved per liter of solution.
- Molality is the ratio of the number of moles of solute dissolved in 1 kg of solvent.
- The number of moles of solute does not change during a dilution.

- 31. MAIN IDEA** Compare and contrast five quantitative ways to describe the composition of solutions.
- 32. Explain** the similarities and differences between a 1M solution of NaOH and a 1m solution of NaOH.
- 33. Calculate** A can of chicken broth contains 450 mg of sodium chloride in 240.0 g of broth. What is the percent by mass of sodium chloride in the broth?
- 34. Solve** How much ammonium chloride ( $\text{NH}_4\text{Cl}$ ), in grams, is needed to produce 2.5 L of a 0.5M aqueous solution?
- 35. Outline** the laboratory procedure for preparing a specific volume of a dilute solution from a concentrated stock solution.

### SECTION 2 REVIEW

- 31.** Molarity, molality, and mole fraction are based on moles of solute per some other quantity; percent by volume and molarity are defined on a per volume of solution basis; molality and mole fraction are based on a per quantity of solvent basis; percent by mass and percent by volume are the only ratios involving percentages.
- 32.** Both solutions contain NaOH (solute) dissolved in water (solvent). The 1m solution contains 1 mole of NaOH per kilogram of water; the 1M solution contains 1 mole of NaOH per liter of solution.
- 33.** 0.19%
- 34.** 70 g  $\text{NH}_4\text{Cl}$
- 35.** Calculate the volume of stock solution needed and add it to a volumetric flask. Add water up to the flask's calibration line.

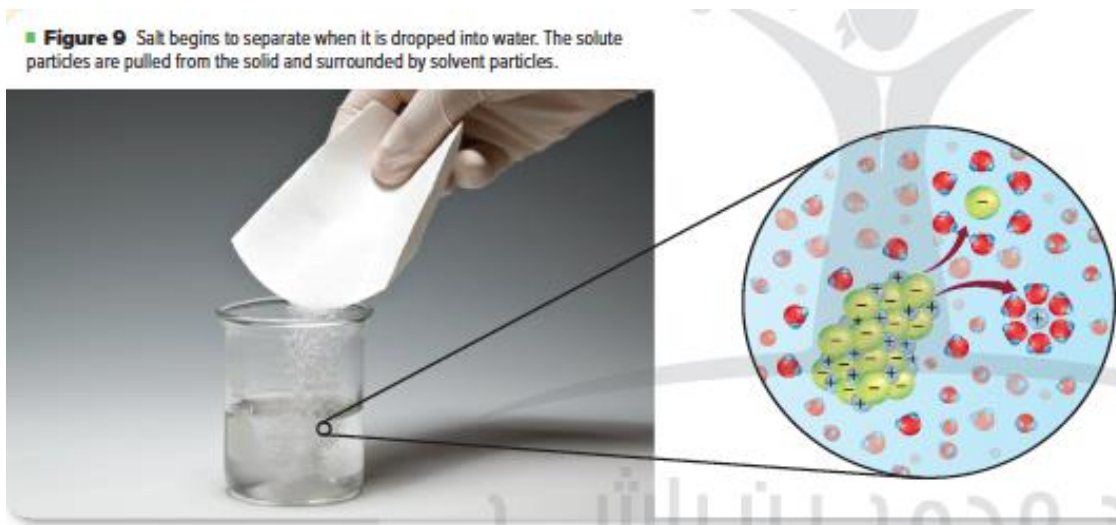
## Section 3 Factors affecting Solvation

### Students Learning outcomes:

1. Explore the effects of changes in the different reaction parameters on the solubility of solid, liquid and gas substances in different solvents.
2. Differentiate qualitatively and quantitatively between the different types of mixtures and solutions.
3. Compare the formation process of solutions that are produced by dissolving ionic and molecular compounds in water, and of solutions that are produced by dissolving non-polar solutes in nonpolar solvents.
4. Describe dynamic and chemical equilibrium while determining the reaction type, equilibrium constants and composition of the reaction medium during the progress of a chemical reaction.
5. Describe the energy changes resulting from the physical changes (e.g.boiling water), chemical reactions (steam cleaning) and nuclear reactions (e.g. nuclear fission, nuclear fusion), in terms of energy releasing or absorbing.

### New vocabulary-

- **Exothermic:** a chemical reaction in which more energy is released than is required to break bonds in the initial reactants.
- **Solvation:** The process of surrounding solute particles with solvent particles to form a solution.



- **Heat of solution:** The overall energy change that occurs during the solution formation process.

- **Unsaturated solution:** is one that contains less dissolved solute for a given temperature and pressure than a saturated solution.
- **Saturated solution:** it contains the maximum amount of dissolved solute for a given amount of solvent at a specific temperature and pressure.
- **Supersaturated solution:** contains more dissolved solute than a saturated solution at the same temperature.
- **Henry's law:** states that at a given temperature, the solubility (S) of a gas in a liquid is directly proportional to the pressure (P) of the gas above the liquid.

#### Henry's Law

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

S represents solubility.

P represents pressure.

At a given temperature, the quotient of solubility of a gas and its pressure is constant.



## Worked Examples –

### EXAMPLE 5

**HENRY'S LAW** If 0.85 g of a gas at 4.0 atm of pressure dissolves in 1.0 L of water at 25°C, how much will dissolve in 1.0 L of water at 1.0 atm of pressure and the same temperature?

#### 1 ANALYZE THE PROBLEM

You are given the solubility of a gas at an initial pressure. The temperature of the gas remains constant as the pressure changes. Because decreasing pressure reduces a gas's solubility, less gas should dissolve at the lower pressure.

**Known**

$$S_1 = 0.85 \text{ g/L}$$

$$P_1 = 4.0 \text{ atm}$$

$$P_2 = 1.0 \text{ atm}$$

**Unknown**

$$S_2 = ? \text{ g/L}$$

#### 2 SOLVE FOR THE UNKNOWN

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

State Henry's law.

$$S_2 = S_1 \left( \frac{P_2}{P_1} \right)$$

Solve Henry's law for  $S_2$ .

$$S_2 = \left( \frac{0.85 \text{ g}}{1.0 \text{ L}} \right) \left( \frac{1.0 \text{ atm}}{4.0 \text{ atm}} \right) = 0.21 \text{ g/L}$$

Substitute  $S_1 = 0.85 \text{ g/L}$ ,  $P_1 = 4.0 \text{ atm}$ , and  $P_2 = 1.0 \text{ atm}$ . Multiply and divide numbers and units.

#### 3 EVALUATE THE ANSWER

The solubility decreased as expected. The pressure on the solution was reduced from 4.0 atm to 1.0 atm, so the solubility should be reduced to one-fourth its original value, which it is. The unit g/L is a solubility unit, and there are two significant figures.

### APPLICATIONS

36. If 0.55 g of a gas dissolves in 1.0 L of water at 20.0 kPa of pressure, how much will dissolve at 110.0 kPa of pressure?
37. A gas has a solubility of 0.66 g/L at 10.0 atm of pressure. What is the pressure on a 1.0 L sample that contains 1.5 g of gas?
38. **Challenge** The solubility of a gas at 7.0 atm of pressure is 0.52 g/L. How many grams of the gas would be dissolved per 1.0 L if the pressure increased 40.0 percent?

36. 3.0 g/L  
37. 23 atm  
38. 0.73 g/L

**Question** If 1.2 g of a gas at 3.5 atm of pressure dissolves in 1.0 L of water at 25°C, how much pressure is needed to dissolve 2.4 g of the gas in 1.0 L of water at the same temperature?

**Answer** 7.0 atm

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

$$P_2 = \left( \frac{S_2}{S_1} \right) P_1 = \left( \frac{2.4 \text{ g/L}}{1.2 \text{ g/L}} \right) 3.5 \text{ atm} \\ = 7.0 \text{ atm}$$

## Section 3 Review

### SECTION 3 REVIEW

#### Section Summary

- The process of solvation involves solute particles surrounded by solvent particles.
- Solutions can be unsaturated, saturated, or supersaturated.
- Henry's law states that at a given temperature, the solubility ( $S$ ) of a gas in a liquid is directly proportional to the pressure ( $P$ ) of the gas above the liquid.

**39. MAINIDEA** Describe factors that affect the formation of solutions.

**40. Define** solubility.

**41. Describe** how intermolecular forces affect solvation.

**42. Explain** on a particle basis why the vapor pressure of a solution is lower than that of the pure solvent.

**43. Summarize** If a seed crystal is added to a supersaturated solution, how would you characterize the resulting solution?

**44. Make and Use Graphs** Use the information in **Table 4** to graph the solubilities of aluminum sulfate, lithium sulfate, and potassium chloride at 0°C, 20°C, 60°C, and 100°C. Which substance's solubility is most affected by increasing temperature?

### SECTION 3 REVIEW

39. Surface area, temperature, and pressure affect the formation of solutions.
40. Solubility refers to the maximum amount of solute that can dissolve in a given amount of solvent at a particular temperature and pressure.
41. The attractive forces between solute and solvent particles overcome the forces holding the solute particles together, thus pulling the solute particles apart.
42. When a solvent contains a solute, fewer solvent particles occupy the surface. Fewer particles escape into the gaseous state.
43. After the excess solute particles crystallize out of solution, the solution is saturated.
44. Refer to the Solution Manual for the graph. Aluminum sulfate shows the greatest change in solubility over the temperature range.

#### SECTION 2 Solution Concentration

**MAINIDEA** Concentration can be expressed in terms of percent or in terms of moles.

- Concentrations can be measured qualitatively and quantitatively.
- Molarity is the number of moles of solute dissolved per liter of solution.

$$\text{molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

- Molality is the ratio of the number of moles of solute dissolved in 1 kg of solvent.

$$\text{molality (m)} = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$$

- The number of moles of solute does not change during a dilution.

$$M_1V_1 = M_2V_2$$

#### VOCABULARY

- concentration
- molarity
- molality
- mole fraction

#### SECTION 3 Factors Affecting Solvation

**MAINIDEA** Factors such as temperature, pressure, and polarity affect the formation of solutions.

- The process of solvation involves solute particles surrounded by solvent particles.
- Solutions can be unsaturated, saturated, or supersaturated.
- Henry's law states that at a given temperature, the solubility ( $S$ ) of a gas in a liquid is directly proportional to the pressure ( $P$ ) of the gas above the liquid.

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

#### VOCABULARY

- solvation
- heat of solution
- unsaturated solution
- saturated solution
- supersaturated solution
- Henry's law

## Chapter 10 Review and MS

### SECTION 1

#### Mastering Concepts

- 54. Explain what is meant by the statement "not all mixtures are solutions."
- 55. What is the difference between a solute and a solvent?
- 56. What is a suspension, and how does it differ from a colloid?
- 57. How can the Tyndall effect be used to distinguish between a colloid and a solution? Why?
- 58. Name a colloid formed from a gas dispersed in a liquid.



Figure 24

- 59. **Salad dressing** What type of heterogeneous mixture is shown in Figure 24? What characteristic is most useful in classifying the mixture?
- 60. What causes the Brownian motion observed in liquid colloids?
- 61. Aerosol sprays are categorized as colloids. Identify the phases of an aerosol spray.

- 54. Solutions are homogeneous mixtures that are uniform in composition with a single phase. Mixtures can also be heterogeneous, where the substances that make them up remain distinct.
- 55. A solute is the substance being dissolved. The solvent is the substance in which the solute dissolves.
- 56. A suspension is a heterogeneous mixture that settles out if left undisturbed. The particles dispersed in a colloid are much smaller than those in a suspension and do not settle out.
- 57. A beam of light is visible in a colloid but not in a solution. Dispersed colloid particles are large enough to scatter light (Tyndall effect).
- 58. Students' answers might include whipped cream or beaten egg whites.
- 59. The mixture is a suspension. Left undisturbed, the mixture components settle out.
- 60. The random particle movements in liquid colloids result from collisions between particles in the mixture.
- 61. The most abundant mixture component is in the gas phase. The dispersed particles are in the liquid phase.

## SECTION 2

### Mastering Concepts

62. What is the difference between percent by mass and percent by volume?
63. What is the difference between molarity and molality?
64. What factors must be considered when creating a dilute solution from a stock solution?
65. How do 0.5M and 2.0M aqueous solutions of NaCl differ?
66. Under what conditions might a chemist describe a solution in terms of molality? Why?

### Mastering Problems

67. According to lab procedure, you stir 25.0 g of  $\text{MgCl}_2$  into 550 mL of water. What is the percent by mass of  $\text{MgCl}_2$  in the solution?
68. How many grams of LiCl are in 275 g of a 15% aqueous solution of LiCl?
69. You need to make a large quantity of a 5% solution of HCl but have only 25 mL HCl. What volume of 5% solution can be made from this volume of HCl?
70. Calculate the percent by volume of a solution created by adding 75 mL of acetic acid to 725 mL of water.
71. Calculate the molarity of a solution that contains 15.7 g of  $\text{CaCO}_3$  dissolved in enough water to make 275 mL of solution.
72. What is the volume of a 3.00M solution made with 122 g of LiF?
73. How many moles of BaS would be used to make  $1.5 \times 10^3$  mL of a 10.0M solution?
74. How much  $\text{CaCl}_2$ , in grams, is needed to make 2.0 L of a 3.5M solution?
75. Stock solutions of HCl with various molarities are frequently prepared. Complete **Table 7** by calculating the volume of concentrated, or 12M, hydrochloric acid that should be used to make 1.0 L of HCl solution with each molarity listed.

**Table 7** HCl Solutions

Molarity of HCl Desired	Volume of 12M HCl Stock Solution Needed (mL)
0.50	
1.0	
1.5	
2.0	
5.0	

62. Percent by mass is a comparison between the mass of solute and the total mass of the solution. Percent by volume is a comparison between the volume of the solute and the total volume of the solution.

63. Molarity is solution concentration expressed as the moles of solute per volume of solution. Molality expresses concentration as moles of solute per kilogram of solvent. Molality does not depend upon the temperature of the solution.

64. The molarity and volume of both stock solution and dilute solution are required in the formula  $M_1V_1 = M_2V_2$ .

65. The 2M solution contains more moles of NaCl per volume of water than the 0.5M solution.

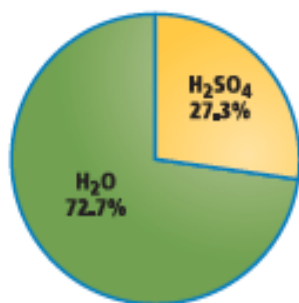
66. Under conditions of changing temperature. Because molality is based on mass, it does not change with temperature.

### Mastering Problems

67. 4.3%
68. 41 g
69. 500 mL



76. How much 5.0M nitric acid ( $\text{HNO}_3$ ), in milliliters, is needed to make 225 mL of 1.0M  $\text{HNO}_3$ ?
77. **Experiment** In the lab, you dilute 55 mL of a 4.0M solution to make 250 mL of solution. Calculate the molarity of the new solution.
78. How many milliliters of 3.0M phosphoric acid ( $\text{H}_3\text{PO}_4$ ) can be made from 95 mL of a 5.0M  $\text{H}_3\text{PO}_4$  solution?
79. If you dilute 20.0 mL of a 3.5M solution to make 100.0 mL of solution, what is the molarity of the dilute solution?
80. What is the molality of a solution that contains 75.3 g of KCl dissolved in 95.0 g of water?
81. How many grams of  $\text{Na}_2\text{CO}_3$  must be dissolved into 155 g of water to create a solution with a molality of 8.20 mol/kg?
82. What is the molality of a solution containing 30.0 g of naphthalene ( $\text{C}_{10}\text{H}_8$ ) dissolved in 500.0 g of toluene?
83. What are the molality and mole fraction of solute in a 35.5 percent by mass aqueous solution of formic acid ( $\text{HCOOH}$ )?



■ Figure 25

84. What is the mole fraction of  $\text{H}_2\text{SO}_4$  in a solution containing the percentage of sulfuric acid and water shown in **Figure 25**?
85. Calculate the mole fraction of  $\text{MgCl}_2$  in a solution created by dissolving 132.1 g of  $\text{MgCl}_2$  in 175 mL of water.

70. 9.4%
71. 0.571M
72. 1.57 L
73. 15 mol
74. 770 g
75. 42 mL; 83 mL; 130 mL; 170 mL; 420 mL
76. 45 mL
77. 0.88M
78. 160 mL
79. 0.70M
80. 10.6 mol/kg
81. 135 g
82. 0.468m
83. 12.0m; 0.177
84. 0.0650
85. 0.125



## SECTION 3

### Mastering Concepts

86. Describe the process of solvation.
87. What are three ways to increase the rate of solvation?
88. Explain the difference between saturated and unsaturated solutions.

### Mastering Problems

89. At a pressure of 1.5 atm, the solubility of a gas is 0.54 g/L. Calculate the solubility when the pressure is doubled.
90. At 4.5 atm of pressure, the solubility of a gas is 9.5 g/L. How much gas, in grams, will dissolve in 1 L if the pressure is reduced by 3.5 atm?

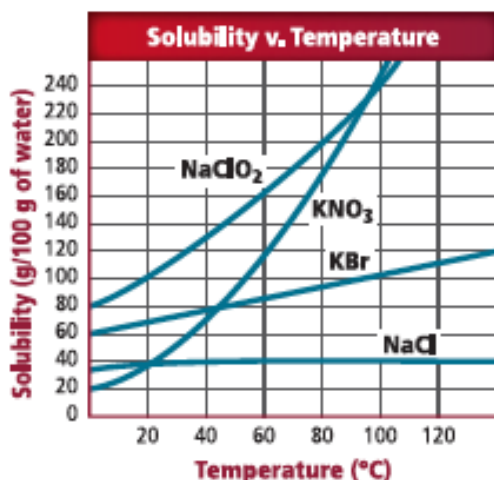


Figure 26

91. Using Figure 26, compare the solubility of potassium bromide (KBr) and potassium nitrate (KNO<sub>3</sub>) at 80°C.
92. The solubility of a gas at 37.0 kPa is 1.80 g/L. At what pressure will the solubility reach 9.00 g/L?
93. Use Henry's law to complete Table 8.

Table 8 Solubility and Pressure	
Solubility (g/L)	Pressure (kPa)
2.9	?
3.7	32
?	39

94. **Soft Drinks** The partial pressure of CO<sub>2</sub> inside a bottle of soft drink is 4.0 atm at 25°C. The solubility of CO<sub>2</sub> is 0.12 mol/L. When the bottle is opened, the partial pressure drops to  $3.0 \times 10^{-4}$  atm. What is the solubility of CO<sub>2</sub> in the open drink? Express your answer in grams per liter.

86. A solute introduced into a solvent is surrounded by solvent particles. Due to the attraction between solute and solvent particles, solute particles are pulled apart and surrounded by solvent particles. Once separated, solute particles disperse into solution.
87. increase the temperature of the solvent, increase the surface area of the solute, agitation
88. A saturated solution contains the maximum amount of solute under a given set of conditions. An unsaturated solution contains less than the maximum amount.

### Mastering Problems

89. 1.1 g/L
90. 2.1 g
91. The solubility of KBr is 95 g/100 g H<sub>2</sub>O. The solubility of KNO<sub>3</sub> is nearly twice as high at the same temperature, at nearly 170 g/100 g H<sub>2</sub>O.
92. 185 kPa
93. 25 kPa; 4.5 g/L
94.  $4.0 \times 10^{-4}$  g/L

# Chapter 11 States of Matter

## Section 1 A Model for reaction rates

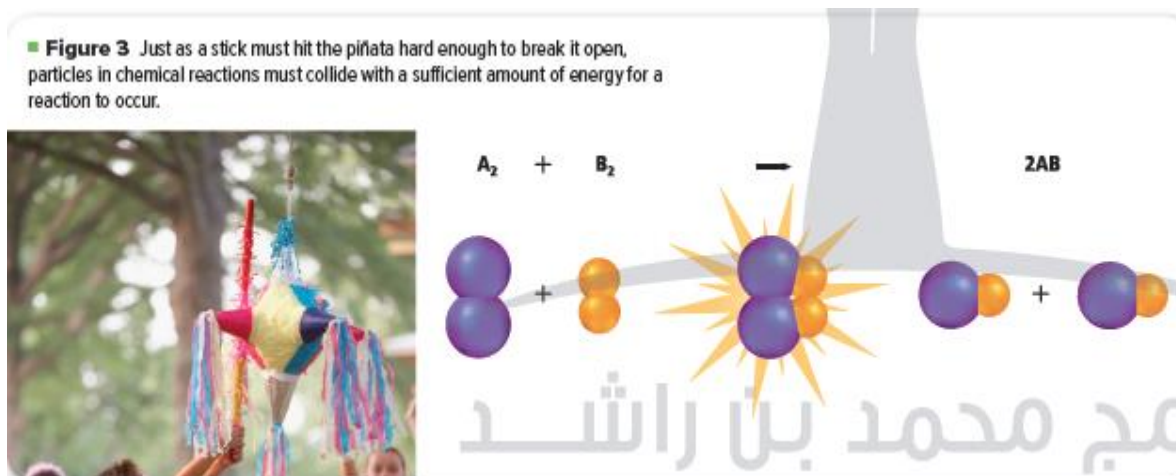
### Student learning outcomes-

- Explain, through reference to a simple chemical reaction how the rate of a reaction is determined by the series of elementary steps that make up the overall reaction mechanism.
- Conclude (by using the collision theory and energy diagrams) the effect of different factors on the chemical reaction rate.
- Conclude the rate laws for chemical reactions.
- Calculate the average and instantaneous reaction rates of simple and complex chemical reactions.
- Interpret the reaction mechanism of a simple or complex reaction while identifying its rate-determining step in several reaction conditions.
- Calculate the average and instantaneous reaction rates of simple and complex chemical reactions.
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- Interpret the reaction mechanism of a simple or complex reaction while identifying its rate-determining step in several reaction conditions.

### New vocabulary-

- **Energy:** the ability to do work or produce heat; it exists in two basic forms: potential energy and kinetic energy
- **Reaction rate:** is the change in concentration of a reactant or product per unit of time, generally expressed as mol/(L .s).
- **Activated complex:** is a temporary, unstable arrangement of atoms in which old bonds are breaking and new bonds are forming.
- **Activation energy ( $E_a$ ):** The minimum amount of energy that reacting particles must have to form the activated complex and lead to a reaction.

- **Collision theory:** which states that atoms, ions, and molecules must collide in order to react.



**Table 1 Collision Theory Summary**

1. Reacting substances (atoms, ions, or molecules) must collide.
2. Reacting substances must collide in the correct orientation.
3. Reacting substances must collide with sufficient energy to form an activated complex.

### Worked Examples –

Write the equation  $\text{CH}_3\text{CH}_2\text{I} + \text{OH}^- \rightarrow \text{CH}_3\text{CH}_2\text{IOH}^- \rightarrow \text{CH}_3\text{CH}_2\text{OH} + \text{I}^-$  on the board and have students identify the activated complex, the products, and the reactants. **Activated complex is  $\text{CH}_3\text{CH}_2\text{IOH}^-$ .** **Products are  $\text{CH}_3\text{CH}_2\text{OH}$  and  $\text{I}^-$ .** **Reactants are  $\text{CH}_3\text{CH}_2\text{I}$  and  $\text{OH}^-$ .**

For the reaction  $\text{A} + \_2\text{B} \rightarrow \_3\text{C}$  at time =  $\_0.000\text{s}$ ,  $[\text{A}] = \_2.00\text{ mol/L}$ ,  $[\text{B}] = \_4.00\text{ mol/L}$ , and  $[\text{C}] = \_0.00\text{ mol/L}$ . At time =  $\_3.00\text{ minutes}$ , A has dropped to  $0.50\text{ mol/L}$ . Have students calculate and express the rate of the reaction over the 3.00-minute time interval in mol A consumed/(L·min), mol B consumed/(L·min), and mol C produced/(L·min). **Rate =  $\_0.500\text{ mol A/(L·min)}$ ; Rate =  $\_1.00\text{ mol B/(L·min)}$ ; Rate =  $\_1.50\text{ mol C/(L·min)}$**

## EXAMPLE 1

**CALCULATE AVERAGE REACTION RATES** In a reaction between butyl chloride ( $\text{C}_4\text{H}_9\text{Cl}$ ) and water, the concentration of  $\text{C}_4\text{H}_9\text{Cl}$  is  $0.220\text{M}$  at the beginning of the reaction. At  $4.00\text{ s}$ , the concentration of  $\text{C}_4\text{H}_9\text{Cl}$  is  $0.100\text{M}$ . Calculate the average reaction rate over the given time period expressed as moles of  $\text{C}_4\text{H}_9\text{Cl}$  consumed per liter per second.

### 1 ANALYZE THE PROBLEM

You are given the initial and final concentrations of the reactant  $\text{C}_4\text{H}_9\text{Cl}$  and the initial and final times. You can calculate the average reaction rate of the chemical reaction using the change in concentration of butyl chloride in four seconds.

#### Known

$$t_1 = 0.00\text{ s}$$

$$t_2 = 4.00\text{ s}$$

$$[\text{C}_4\text{H}_9\text{Cl}] \text{ at } t_1 = 0.220\text{M}$$

$$[\text{C}_4\text{H}_9\text{Cl}] \text{ at } t_2 = 0.100\text{M}$$

#### Unknown

$$\text{Average reaction rate} = ? \text{ mol}/(\text{L}\cdot\text{s})$$

### 2 SOLVE FOR THE UNKNOWN

$$\begin{aligned}\text{Average reaction rate} &= \frac{[\text{C}_4\text{H}_9\text{Cl}] \text{ at } t_2 - [\text{C}_4\text{H}_9\text{Cl}] \text{ at } t_1}{t_2 - t_1} \\ &= -\frac{0.100\text{M} - 0.220\text{M}}{4.00\text{ s} - 0.00\text{ s}} \\ &= -\frac{0.100\text{ mol/L} - 0.220\text{ mol/L}}{4.00\text{ s} - 0.00\text{ s}}\end{aligned}$$

$$\text{Average reaction rate} = \frac{0.120\text{ mol/L}}{4.00\text{ s}} = 0.0300\text{ mol}/(\text{L}\cdot\text{s})$$

State the average reaction rate equation.

Substitute  $t_2 = 4.00\text{ s}$ ,  $t_1 = 0.00\text{ s}$ ,  $[\text{C}_4\text{H}_9\text{Cl}]$  at  $t_2 = 0.100\text{ M}$ , and  $[\text{C}_4\text{H}_9\text{Cl}]$  at  $t_1 = 0.220\text{M}$ .

Substitute mol/L for  $M$  and perform the calculations. Note that the minus sign cancels out.

### 3 EVALUATE THE ANSWER

The average reaction rate of  $0.0300$  moles  $\text{C}_4\text{H}_9\text{Cl}$  consumed per liter per second is reasonable based on the starting and ending amounts. The answer is correctly expressed in three significant figures.

## APPLICATIONS

Use the data in the following table to calculate the average reaction rates.

Experimental Data for $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$			
Time (s)	$[\text{H}_2]$ (M)	$[\text{Cl}_2]$ (M)	$[\text{HCl}]$ (M)
0.00	0.030	0.050	0.000
4.00	0.020	0.040	

1.  $0.0025\text{ mol}/(\text{L}\cdot\text{s})$
2.  $0.0025\text{ mol}/(\text{L}\cdot\text{s})$
3.  $0.020\text{M}$

1. Calculate the average reaction rate expressed in moles  $\text{H}_2$  consumed per liter per second.
2. Calculate the average reaction rate expressed in moles  $\text{Cl}_2$  consumed per liter per second.
3. **Challenge** If the average reaction rate for the reaction, expressed in moles of  $\text{HCl}$  formed, is  $0.0050\text{ mol}/\text{L}\cdot\text{s}$ , what concentration of  $\text{HCl}$  would be present after  $4.00\text{ s}$ ?

## Section 1 Review

### SECTION 1 REVIEW

#### Section Summary

- The rate of a chemical reaction is expressed as the rate at which a reactant is consumed or the rate at which a product is formed.
  - Reaction rates are generally calculated and expressed in moles per liter per second ( $\text{mol}/(\text{L} \cdot \text{s})$ ).
  - In order to react, the particles in a chemical reaction must collide.
4. **MAIN IDEA** **Relate** collision theory to reaction rate.
  5. **Explain** what the reaction rate indicates about a particular chemical reaction.
  6. **Compare** the concentrations of the reactants and products during the course of a chemical reaction (assuming no additional reactants are added).
  7. **Compare** the average reaction rate measured over an initial, short time interval to one measured over a long time interval.
  8. **Describe** the relationship between activation energy and the rate of a reaction.
  9. **Summarize** what happens during the brief existence of an activated complex.
  10. **Apply** collision theory to explain why collisions between two reacting particles do not always result in the formation of a product.
  11. **Interpret** how the speed of a chemical reaction is related to the spontaneity of the reaction.
  12. **Calculate** the average rate of a reaction between hypothetical molecules A and B if the concentration of A changes from  $1.00M$  to  $0.50M$  in  $2.00\text{ s}$ .

### SECTION 1 REVIEW

4. In order for a reaction to occur, molecules, atoms, or ions must collide. The frequency, orientation, and energy of those collisions determine the rate of the overall reaction.
5. The reaction rate indicates the rate of change of the concentration of a reactant or product in  $\text{mol}/(\text{L} \cdot \text{s})$ .
6. The concentrations of the reactants decrease, and the concentrations of the products increase at the same rate.
7. As the concentration of the reactants decreases, the reaction rate decreases. Thus, the longer the time period, the smaller the average reaction rate.
8. The higher the activation energy, the slower the rate of the reaction.
9. Bonds in the reactants are in the process of breaking, while new bonds are beginning to form to produce the products.
10. The collision must be in a correct orientation and have sufficient energy to form the activated complex.
11. The speed and spontaneity of a chemical reaction are not related.
12.  $\text{Rate} = 0.25\text{ mol}/(\text{L} \cdot \text{s})$



## Section 3 Reaction Rate Law

### Student learning outcomes-

- Explain, through reference to a simple chemical reaction, how the rate of a reaction is determined by the series of elementary steps that make up the overall reaction mechanism.
- Relate the order of a reaction to the rate law for this reaction while employing the method of initial rates in determining the rate, concentration, reaction order and specific rate constant of a chemical reaction .
- Employ the rate law to determine the rate, concentration, reaction order, reactants orders and specific rate constant of a chemical reaction .
- Calculate the average and instantaneous reaction rates of simple and complex chemical reactions.

### New vocabulary-

- **Reactant:** the starting substance in a chemical reaction New Vocabulary rate law specific rate constant reaction order method of initial rates.
- **Rate Law:** expresses the relationship between the rate of a chemical reaction and the concentration of reactants at a given temperature.

**One-Step Reaction Rate Law**

$\text{rate} = k[A]$  [A] represents the concentration of a reactant;  
k is a constant.

The rate of a one-step reaction is the product of the concentration of the reactant and a constant.

- **Specific rate constant:** a numerical value that relates the reaction rate and the concentrations of reactants at a given temperature.
- **Reaction order:** defines how the rate is affected by the concentration of that reactant.
- **Method of initial rates** determines reaction order by comparing the initial rates of a reaction carried out with varying reactant concentrations.

## Worked Examples –

### APPLICATIONS

- 19.** Write the rate law for the reaction  $aA \rightarrow bB$  if the reaction is third order in A. [B] is not part of the rate law.
- 20.** The rate law for the reaction  $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$  is first order in  $O_2$  and third order overall. What is the rate law for the reaction?
- 21.** Given the experimental data below, use the method of initial rates to determine the rate law for the reaction  $aA + bB \rightarrow \text{products}$ .  
(Hint: Any number to the zero power equals one. For example,  $(0.22)^0 = 1$  and  $(55.6)^0 = 1$ .)

**Practice Problem 21 Experimental Data**

Trial	Initial [A](M)	Initial [B](M)	Initial Rate (mol/(L·s))
1	0.100	0.100	$2.00 \times 10^{-3}$
2	0.200	0.100	$2.00 \times 10^{-3}$
3	0.200	0.200	$4.00 \times 10^{-3}$

- 22. Challenge** The rate law for the reaction  $CH_3CHO(g) \rightarrow CH_4(g) + CO(g)$  is  $\text{Rate} = k[CH_3CHO]^2$ . Use this information to fill in the missing experimental data below.

**Practice Problem 22 Experimental Data**

Trial	Initial $[CH_3CHO](M)$	Initial Rate (mol/(L·s))
1	$2.00 \times 10^{-3}$	$2.70 \times 10^{-11}$
2	$4.00 \times 10^{-3}$	$10.8 \times 10^{-11}$
3	$8.00 \times 10^{-3}$	

- 19.**  $\text{Rate} = k[A]^3$
- 20.**  $\text{Rate} = k[O_2][NO]^2$
- 21.** Examining trials 1 and 2, doubling [A] has no effect on the rate; therefore, the reaction is zero order in A. Examining trials 2 and 3, doubling [B] doubles the rate; therefore, the reaction is first order in B.  $\text{Rate} = k[A]^0[B] = k[B]$
- 22.** Examining trials 1 and 2, doubling  $[CH_3CHO]$  increases the rate by a factor of four. Examining trials 2 and 3,  $[CH_3CHO]$  is again doubled so the rate again must increase by a factor of four. Therefore, the rate for trial 3 is  $43.2 \times 10^{-11} \text{ mol/(L·s)}$ .

## Section 3 Review

### SECTION 3 REVIEW

#### Section Summary

- The mathematical relationship between the rate of a chemical reaction at a given temperature and the concentrations of reactants is called the rate law.
- The rate law for a chemical reaction is determined experimentally using the method of initial rates.

- 23. MAIN IDEA** Explain what the rate law for a chemical reaction tells you about the reaction.
- 24. Apply** the rate-law equations to show the difference between a first-order reaction with a single reactant and a second-order reaction with a single reactant.
- 25. Explain** the function of the specific rate constant in a rate-law equation.
- 26. Explain** Under what circumstance is the specific rate constant ( $k$ ) not a constant? What does the size of  $k$  indicate about the rate of a reaction?
- 27. Suggest** a reason why, when given the rate of a chemical reaction, it is important to know that the reaction rate is an average reaction rate.
- 28. Explain** how the exponents in the rate equation for a chemical reaction relate to the coefficients in the chemical equation.
- 29. Determine** the overall reaction order for a reaction between A and B for which the rate law is  $\text{rate} = k[A]^2[B]^2$ .
- 30. Design an Experiment** Explain how you would design an experiment to determine the rate law for the general reaction  $aA + bB \rightarrow \text{products}$  using the method of initial rates.

### SECTION 3 REVIEW

- 23.** The rate law expresses the mathematical relationship between the reaction's rate and the concentrations of reactants.
- 24.** First order reaction:  $\text{Rate} = k[A]$ ; Second-order reaction:  $\text{Rate} = k[A]^2$
- 25.** The specific rate constant ( $k$ ) relates reaction rate and concentration at a specific temperature.
- 26.**  $k$  changes with temperature. The larger the value of  $k$  the faster the reaction.
- 27.** The rate of a reaction decreases over time as reactant concentrations decrease. Therefore, the rate is an average over time rather than the rate at a given instant.
- 28.** There is no relationship in general. In the relatively rare case of a single-step reaction with a single activated complex, the exponents are equal to the coefficients.
- 29.** The overall reaction is fourth.
- 30.** Determine the order of reactant A by measuring the reaction rate for several trials in which [A] is varied while [B] remains constant. Determine the order of reactant B by measuring the reaction rate for several trials in which [B] is varied while [A] remains constant.

## Chapter 11 Review and MS

### SECTION 1

#### Mastering Concepts

40. What happens to the concentrations of the reactants and products during the course of a chemical reaction?
41. Explain what is meant by the average rate of a reaction.
42. How would you express the rate of the chemical reaction  $A \rightarrow B$  based on the concentration of Reactant A? How would that rate compare with the reaction rate based on the Product B?
43. What is the role of the activated complex in a chemical reaction?
44. Suppose two molecules that can react collide. Under what circumstances do the colliding molecules not react?

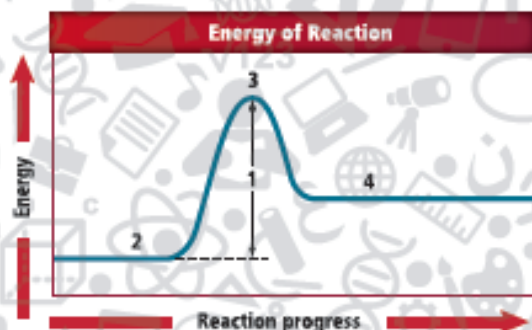


Figure 21

45. Figure 21 is an energy diagram for a reaction. Match the appropriate number with the quantity it represents.
  - a. reactants
  - b. activated complex
  - c. products
  - d. activation energy
46. If  $A \rightarrow B$  is exothermic, how does the activation energy for the forward reaction compare with the activation energy for the reverse reaction ( $A \leftarrow B$ )?

#### Mastering Problems

47. In the gas-phase reaction,  $I_2 + Cl_2 \rightarrow 2ICl$ ,  $[I_2]$  changes from 0.400M at 0.00 min to 0.300M at 4.00 min. Calculate the average reaction rate in moles of  $I_2$  consumed per liter per minute.
48. In a reaction  $Mg(s) + 2HCl(aq) \rightarrow H_2(g) + MgCl_2(aq)$ , 6.00 g of Mg was present at 0.00 min. After 3.00 min, 4.50 g of Mg remained. Express the average rate as mol Mg consumed/min.
49. If a chemical reaction occurs at the rate of  $2.25 \times 10^{-2}$  moles per liter per second at 322 K, what is the rate expressed in moles per liter per minute?

40. The concentration of reactants decreases and the concentration of products increases.
41. The average rate of a reaction is the change in the concentration of a reactant or product that occurs during a specified time interval.
42. The rate may be expressed as the decrease in  $[A]$  per unit time:  
 $Rate = \Delta[A]/\Delta t$ . Numerically, the two rates would be equal; however,  $\Delta[A]$  would be negative, showing a decrease in concentration of the reactant as the reaction proceeds, and  $\Delta[B]$  would be positive, showing an increase in concentration of the product as the reaction proceeds.
43. The activated complex is a transition state between reactants and products
44. The molecules do not react if they collide with insufficient energy or lack favorable orientations at the instant of impact.
45. a. reactants 2  
b. activated complex 3  
c. products 4  
d. activation energy 1
46. The activation energy for the forward reaction is less than the activation energy for the reverse reaction.

#### Mastering Problems

47. The average reaction rate is 0.0250 mol  $I_2$  consumed per liter per minute.
48. Average rate =  $0.0617 \text{ mol}/3.00 \text{ min} = 2.06 \times 10^{-2} \text{ mol/min}$
49. Rate =  $3.75 \times 10^{-4} \text{ mol}/(\text{L}\cdot\text{min})$



## SECTION 3

### Mastering Concepts

60. In the method of initial rates used to determine the rate law for a chemical reaction, what is the significance of the word *initial*?
61. Why must the rate law for a chemical reaction be based on experimental evidence rather than the balanced equation for the reaction?
62. Assume that the rate law for a generic chemical reaction is  $\text{rate} = [A][B]^3$ . What is the reaction order in A, the reaction order in B, and the overall reaction order?
63. Consider the generic chemical reaction:  $A + B \rightarrow AB$ . Based on experimental data, the reaction is second order in Reactant A. If the concentration of A is halved, and all other conditions remain unchanged, how does the reaction rate change?

### Mastering Problems

64. The instantaneous rate data in Table 3 were obtained for the reaction  $\text{H}_2(\text{g}) + 2\text{NO}(\text{g}) \rightarrow \text{H}_2\text{O}(\text{g}) + \text{N}_2\text{O}(\text{g})$  at a given temperature and concentration of NO. How does the instantaneous rate of this reaction change as the initial concentration of  $\text{H}_2$  is changed? Based on the data, is  $[\text{H}_2]$  part of the rate law? Explain.

**Table 3** Reaction Between  $\text{H}_2(\text{g})$  and  $\text{NO}(\text{g})$

$[\text{H}_2]$ (mol/L)	Instantaneous Rate (mol/L·s)
0.18	$6.00 \times 10^{-3}$
0.32	$1.07 \times 10^{-2}$
0.58	$1.93 \times 10^{-2}$

65. Suppose that a generic chemical reaction has the rate law of  $\text{rate} = [A]^2[B]^3$  and that the reaction rate under a given set of conditions is  $4.5 \times 10^{-4} \text{ mol/(L} \cdot \text{min)}$ . If the concentrations of both A and B are doubled and all other reaction conditions remain constant, how will the reaction rate change?
66. The experimental data in Table 4 were obtained for the decomposition of azomethane ( $\text{CH}_3\text{N}_2\text{CH}_3$ ) at a particular temperature according to the equation  $\text{CH}_3\text{N}_2\text{CH}_3(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g}) + \text{N}_2(\text{g})$ . Use the data to determine the reaction's experimental rate law.

**Table 4** Decomposition of Azomethane

Experiment Number	Initial $[\text{CH}_3\text{N}_2\text{CH}_3]$	Initial Reaction Rate
1	0.012M	$2.5 \times 10^{-6} \text{ mol/(L} \cdot \text{s)}$
2	0.024M	$5.0 \times 10^{-6} \text{ mol/(L} \cdot \text{s)}$

60. The initial rate means the moment at which the reactants are mixed at the stated concentrations; however, the rate begins to decrease the instant the reaction starts.
61. Most chemical reactions occur in more than one step.
62. The reaction is first order in A, third order in B, and fourth order overall.
63. The rate decreases to one-fourth its initial value.

### Mastering Problems

64. As the concentration of  $\text{H}_2$  increases, the reaction rate increases. Yes; the fact that the rate increases as  $[\text{H}_2]$  increases indicates that  $[\text{H}_2]$  appears in the rate law and is raised to a positive exponent. The data show that the ratios between two successive  $[\text{H}_2]$  values are equal to the ratios between the corresponding instantaneous rate values. Therefore, the positive exponent of  $[\text{H}_2]$  in the rate law is one.
65. Because  $(2^2)(2^3) = 32$ , the reaction rate will increase by a factor of 32 to  $1.4 \times 10^{-2} \text{ mol/(L} \cdot \text{min)}$ .
66. In experiment 2, doubling the concentration of azomethane,  $\text{CH}_3\text{N}_2\text{CH}_3$ , doubles the reaction rate. Therefore,  $\text{Rate} = k[\text{CH}_3\text{N}_2\text{CH}_3]$ .

67. Use the data in Table 4 to calculate the value of the specific rate constant,  $k$ .

68. At the same temperature, predict the reaction rate when the initial concentration of  $\text{CH}_3\text{N}_2\text{CH}_3$  is 0.048M. Use the

$$67. k = 2.1 \times 10^{-4} \text{ s}^{-1}$$

$$68. \text{Rate} = 1.0 \times 10^{-5} \text{ mol/(L} \cdot \text{s)}$$