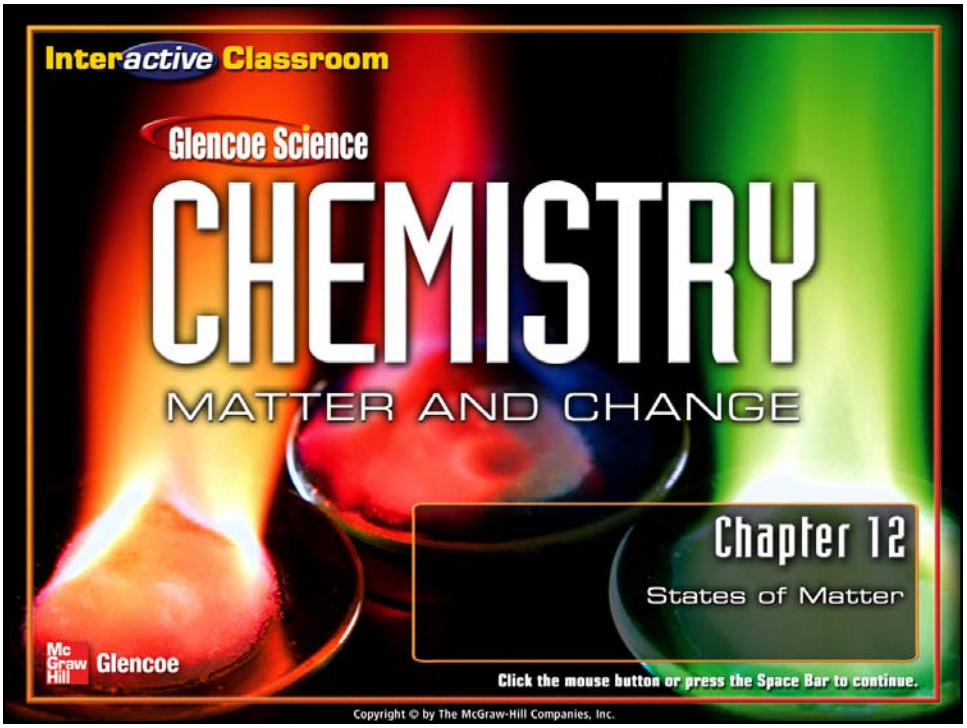
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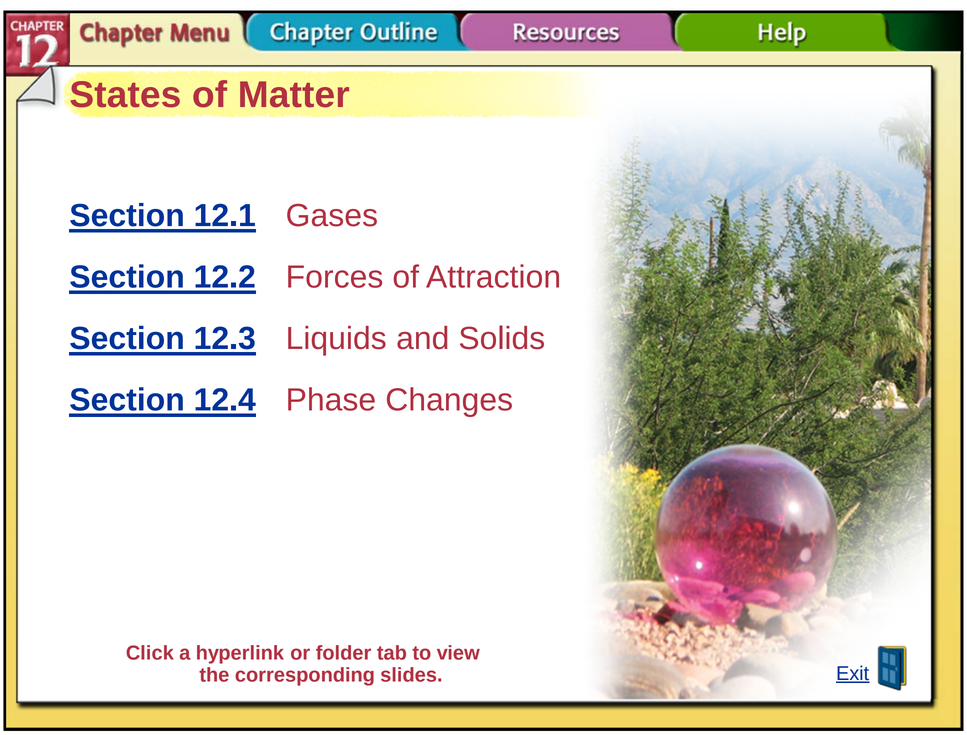
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Section 12.1 Gases

Objectives

- **Use** the kinetic-molecular theory to explain the behavior of gases.
- **Describe** how mass affects the rates of diffusion and effusion.
- **Explain** how gas pressure is measured and calculate the partial pressure of a gas.

Review Vocabulary

kinetic energy: energy due to motion

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Section 12.1 Gases (cont.)

New Vocabulary

- kinetic-molecular theory
- elastic collision
- temperature
- diffusion
- Graham's law of effusion
- pressure
- barometer
- pascal
- atmosphere
- Dalton's law of partial pressures

MAIN Idea **Gases expand, diffuse, exert pressure, and can be compressed because they are in a low density state consisting of tiny, constantly-moving particles.**

•4

States of Matter

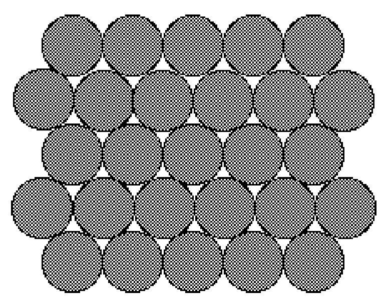
Matter exists in one of four states:

1. Solids
2. Liquids
3. Gases
4. Plasmas

•5

Solids

- matter that has a definite shape and volume.
- have a high density.
- expand only slightly when heated.
- have a shape that does not depend upon the shape of the container into which it is placed.
- are almost incompressible.
- Have particles that exhibit vibrational motion.



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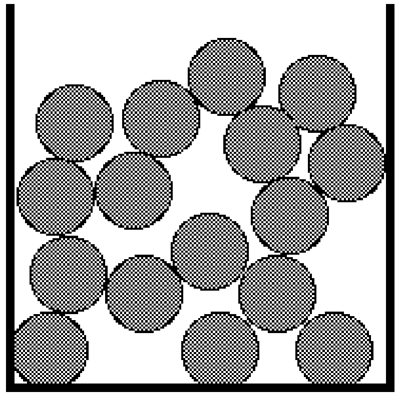
Liquids



- matter that flows, has a fixed volume, and takes the shape of its container.
- are generally less dense than solids.
- expand slightly when heated.
- are almost incompressible.
- Have particles that exhibit both vibrational and rotational motion.

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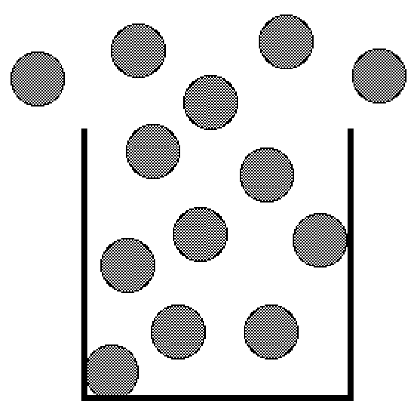
Gases

- matter that takes both the shape and the volume of its container.
- expand without limit to fill any space they are found.
- are easily compressed.
- have particles that exhibit **vibrational**, **rotational**, and **translational** motion.
- **Vapour** and **gas** are two different terms.
- A vapour is a substance that, although it is in a gaseous state, is generally a liquid or a solid at room temperature.



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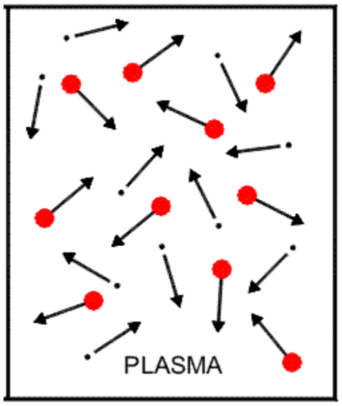
Plasma



- is a gas consisting of charged ions and electrons.
- is like two gases intermixed – a gas of electrons and a gas of ions.
- is made by heating a gas until the nucleus and the electrons around it can no longer stay together.
- The atoms are stripped of their electrons and a plasma is formed.

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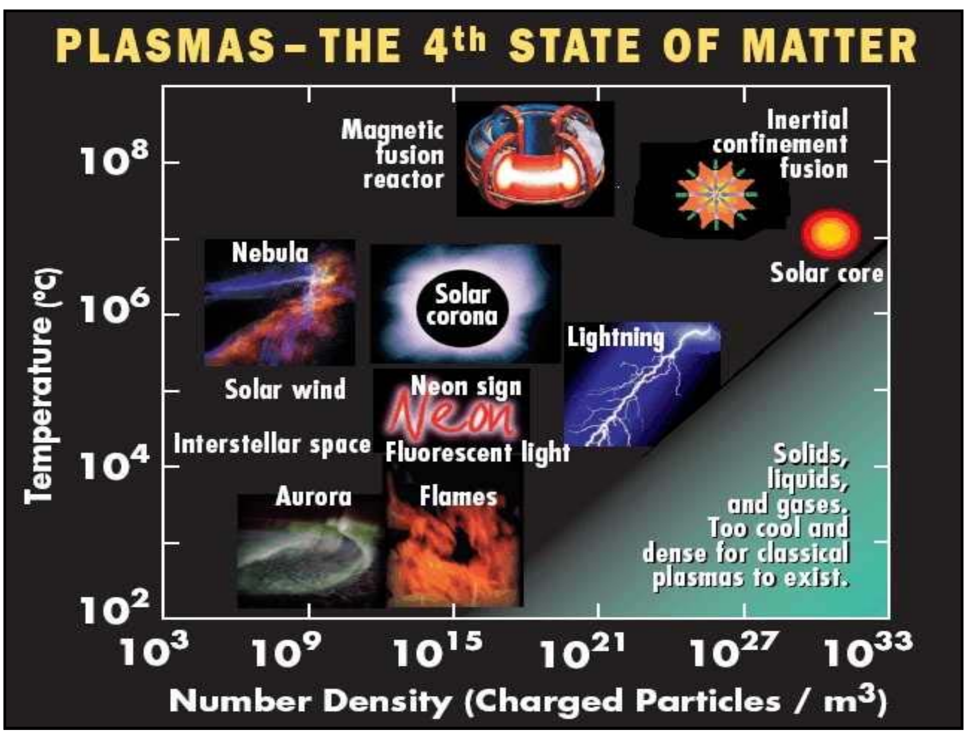

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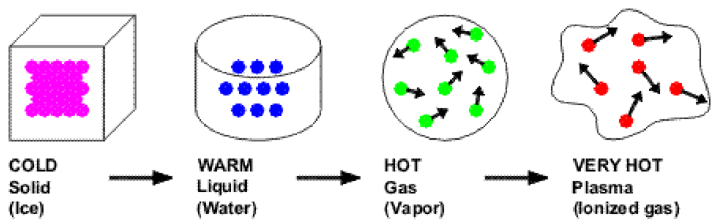
Plasma (cont.)

- found in the stars and the interstellar environment.
- makes up most of our universe (around 99%).
- On Earth, it does not exist in a natural form apart from in lightning and the **Aurora Borealis**.
- most common application - fluorescent light.



The illustration shows how matter changes its state by the addition of thermal energy.

The amount and type of movement is smallest in solids, greater in liquids, and greatest in gases.

The Kinetic-Molecular Theory

- Kinetic-molecular theory explains the different properties of solids, liquids, and gases.
- Atomic composition affects chemical properties.
- Atomic composition also affects physical properties.
- The kinetic-molecular theory describes the behavior of matter in terms of particles in motion.
- Gases consist of small particles separated by empty space.
- Gas particles are too far apart to experience significant attractive or repulsive forces.

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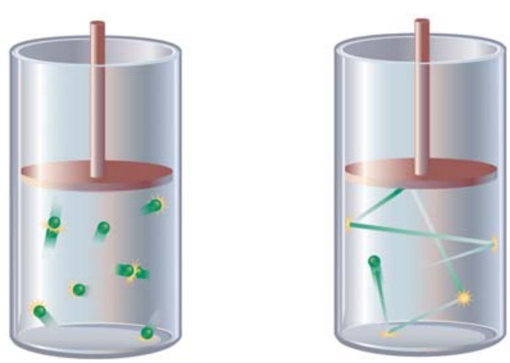
The Kinetic-Molecular Theory (cont.)

- Gas particles are in constant random motion.
- An **elastic collision** is one in which no kinetic energy is lost.



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The Kinetic-Molecular Theory (cont.)

- Kinetic energy of a particle depends on mass and velocity.

$$KE = \frac{1}{2}mv^2$$

- **Temperature** is a measure of the average kinetic energy of the particles in a sample of matter.

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Energy

- is the capacity for doing work.
- Can exist in many forms, chemical, nuclear, electrical, radiant, mechanical and thermal.
- Nearly all chemical and physical changes in nature involve the absorption or emission of energy.

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Energy (cont.)

- Energy may be either potential or kinetic.
- **Potential energy** is energy of position or composition. It is stored energy.
- **Kinetic energy** is energy of motion.
- Potential energy can be converted into kinetic energy.

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Energy (cont.)

- Forms of energy are **interconvertable**.
- Many of the activities we do each day involve a series of energy conversions.
- Consider a bike ride on a sunny summer day.
- The series of energy conversions are a follows ...

Radiant → chemical → mechanical → heat

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Energy (cont.)

- Trace the transfer of energy from a power dam in northern Manitoba to the power drill used in the VCI wood shop . . .

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- The law of conservation of energy states that energy can not be created or destroyed.
- Energy can only be converted from one form to another.
- Production of **heat** is often considered to be waste energy.

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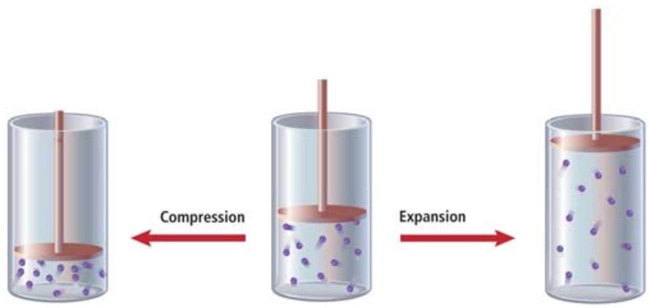
Explaining the Behavior of Gases

- Great amounts of space exist between gas particles.
- Compression reduces the empty spaces between particles.



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Explaining the Behavior of Gases (cont.)

- Gases easily flow past each other because there are no significant forces of attraction.
- **Diffusion** is the movement of one material through another.
- Effusion is a gas escaping through a tiny opening.

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Explaining the Behavior of Gases (cont.)

- **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.

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

- Graham's law also applies to diffusion.

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

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Gas Pressure

- **Pressure** is defined as force per unit area.
- Gas particles exert pressure when they collide with the walls of their container.

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Gas Pressure (cont.)

- The particles in the earth's atmosphere exert pressure in all directions called air pressure.
- There is less air pressure at high altitudes because there are fewer particles present, since the force of gravity is less.

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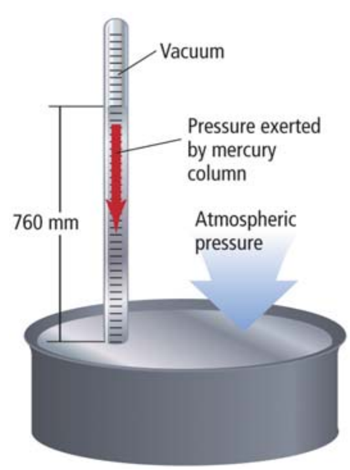
Gas Pressure (cont.)

- Torricelli invented the barometer.
- **Barometers** are instruments used to measure atmospheric air pressure.



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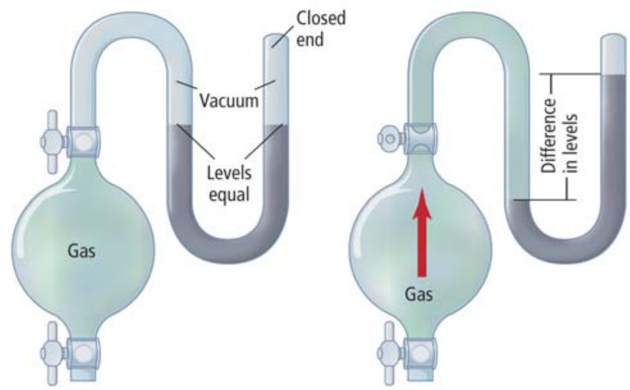
Gas Pressure (cont.)

- Manometers measure gas pressure in a closed container.



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Gas Pressure (cont.)

- The SI unit of force is the newton (N).
- One **pascal**(Pa) is equal to a force of one Newton per square meter or N/m^2.
- One **atmosphere** is equal to 760 mm Hg or 101.3 kilopascals.

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Gas Pressure (cont.)

Table 12.1 Comparison of Pressure Units

Unit	Number Equivalent to 1 atm	Number Equivalent to 1 kPa
Kilopascal (kPa)	101.3 kPa	—
Atmosphere (atm)	—	0.009869 atm
Millimeters of mercury (mm Hg)	760 mm Hg	7.501 mm Hg
Torr	760 torr	7.501 torr
Pounds per square inch (psi or lb/in^2)	14.7 psi	0.145 psi
Bar	1.01 bar	100 kPa

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Gas Pressure (cont.)

- **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 of the mixture.
- The partial pressure of a gas depends on the number of moles, size of the container, and temperature and is independent of the type of gas.

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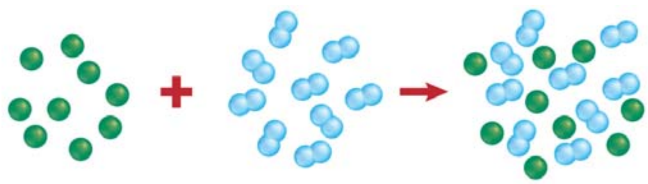
Gas Pressure (cont.)

$$P_{total} = P1 + P2 + P3 + ...Pn$$



- Partial pressure can be used to calculate the amount of gas produced in a chemical reaction.

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Section 12.1 Assessment

CheckPoint

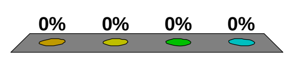
The average of kinetic energy of particles in a substance is measured by its ____.

A. mass

B. density

C. temperature

D. pressure



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Section 12.1 Assessment

CheckPoint

One mole of oxygen in a 5.0 liter container has the same partial pressure as one mol of hydrogen in the same container. This is a demonstration of what law?

A. law of conservation of mass

B. law of definite proportions

C. law of conservation of energy

D. Dalton's law of partial pressures

A	B	C	D
0%	0%	0%	0%

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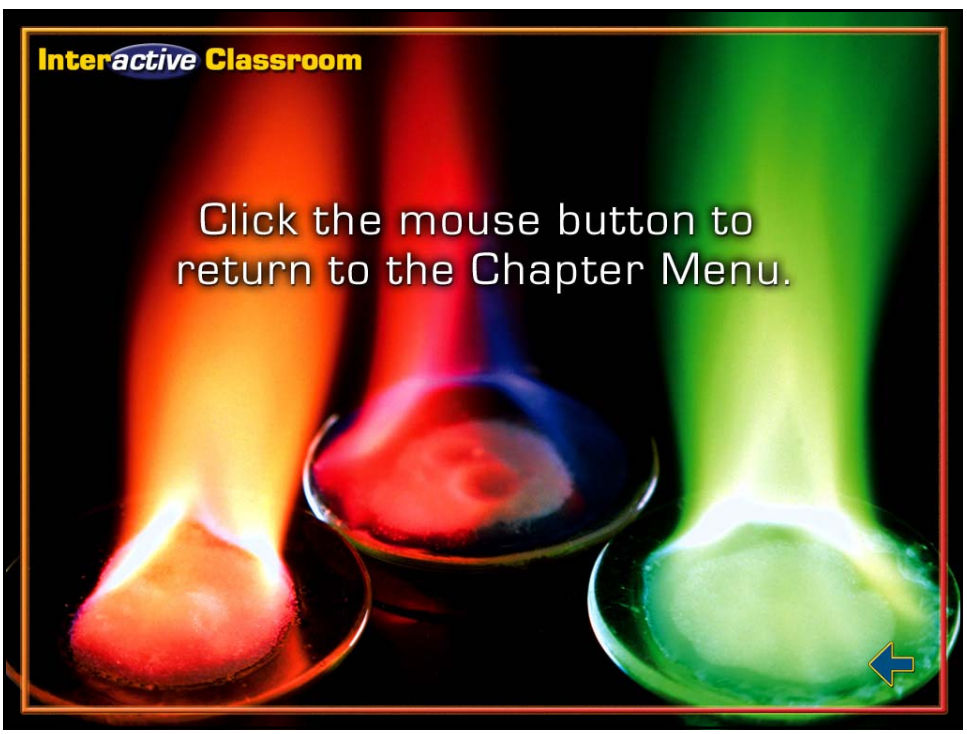

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Section 12.2 Forces of Attraction

Objectives

- **Describe** intramolecular forces.
- **Compare and contrast** intermolecular forces.

Review Vocabulary

polar covalent: a type of bond that forms when electrons are not shared equally

New Vocabulary

dispersion force

dipole-dipole force

hydrogen bond

MAIN Idea

Intermolecular forces—including dispersion forces, dipole-dipole forces, and hydrogen bonds—determine a substance's state at a given temperature.

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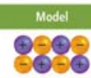
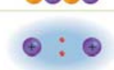
Intermolecular Forces

- Attractive forces between molecules cause some materials to be solids, some to be liquids, and some to be gases at the same temperature.

Table 12.2 Comparison of Intramolecular Forces

Force	Model	Basis of Attraction	Example
Ionic		cations and anions	NaCl
Covalent		positive nuclei and shared electrons	H_2
Metallic		metal cations and mobile electrons	Fe

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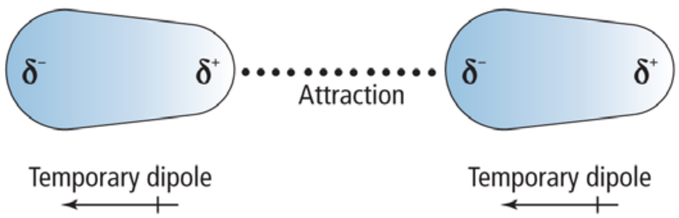
Intermolecular Forces (cont.)

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



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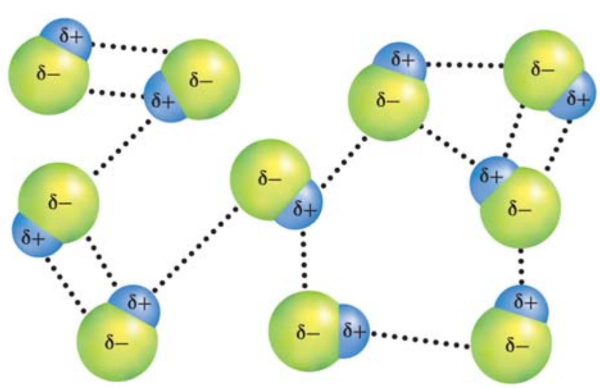
Intermolecular Forces (cont.)

- **Dipole-dipole forces** are attractions between oppositely charged regions of polar molecules.



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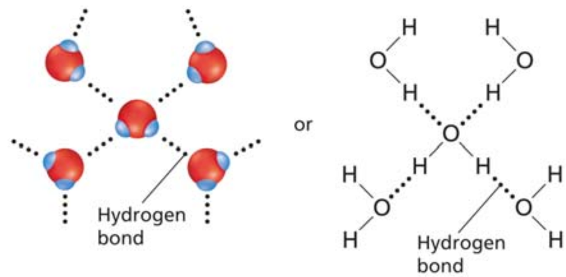
Intermolecular Forces (cont.)

- **Hydrogen bonds** are special dipole-dipole attractions that occur between molecules that contain a hydrogen atom bonded to a small, highly electronegative atom with at least one lone pair of electrons, typically fluorine, oxygen, or nitrogen.



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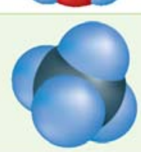
Intermolecular Forces (cont.)

Table 12.3 Properties of Three Molecular Compounds

Compound	Molecular Structure	Molar Mass (g)	Boiling Point (°C)
Water (H_2O)		18.0	100
Methane (CH_4)		16.0	–33.4
Ammonia (NH_3)		17.0	–164

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Section 12.2 Assessment

CheckPoint

A hydrogen bond is a type of ____.

A. dispersion force

B. ionic bond

C. covalent bond

D. dipole-dipole force

0% 0% 0% 0%

A B C D

•41

Section 12.2 Assessment

CheckPoint

Which of the following molecules can form hydrogen bonds?

A. CO_2

B. C_2H_6

C. NH_3

D. H_2

0% 0% 0% 0%

A B C D

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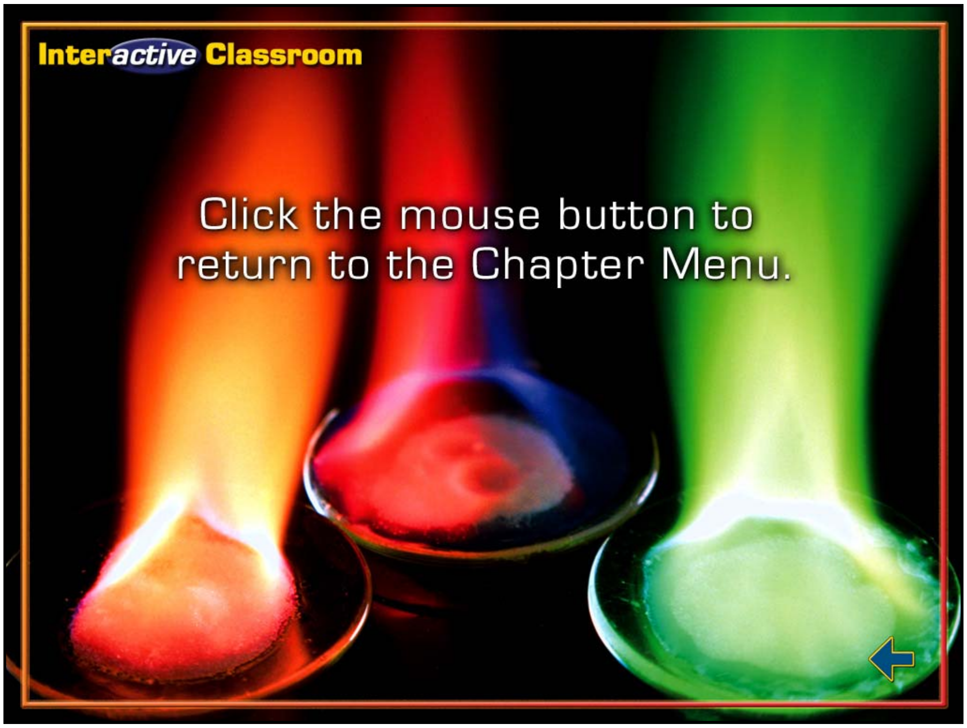

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Section 12.3 Liquids and Solids

Objectives

- **Contrast** the arrangement of particles in liquids and solids.
- **Describe** the factors that affect viscosity.
- **Explain** how the unit cell and crystal lattice are related.

Review Vocabulary

meniscus: the curved surface of a column of liquid

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Section 12.3 Liquids and Solids (cont.)

New Vocabulary

viscosity

surface tension

surfactant

crystalline solid

unit cell

allotrope

amorphous solid

MAIN Idea **The particles in solids and liquids have a limited range of motion and are not easily compressed.**

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Liquids

- Forces of attraction keep molecules closely packed in a fixed volume, but not in a fixed position.
- Liquids are much denser than gases because of the stronger intermolecular forces holding the particles together.
- Large amounts of pressure must be applied to compress liquids to very small amounts.

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Liquids (cont.)

- Fluidity is the ability to flow and diffuse; liquids and gases are fluids.
- <u>Viscosity</u> is a measure of the resistance of a liquid to flow and is determined by the type of intermolecular forces, size and shape of particles, and temperature.

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Liquids (cont.)

- The stronger the intermolecular attractive forces, the higher the viscosity.
- Larger molecules create greater viscosity.
- Long chains of molecules result in a higher viscosity.
- Increasing the temperature increases viscosity because the added energy allows the molecules to overcome intermolecular forces and flow more freely.

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Liquids (cont.)

- **Surface tension** is the energy required to increase the surface area of a liquid by a given amount.
- **Surfactants** are compounds that lower the surface tension of water.

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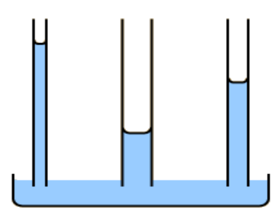
Liquids (cont.)

- Cohesion is the force of attraction between identical molecules.
- Adhesion is the force of attraction between molecules that are different.
- Capillary action is the upward movement of liquid into a narrow cylinder, or capillary tube.



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Solids

- Solids contain particles with strong attractive intermolecular forces.
- Particles in a solid vibrate in a fixed position.
- Most solids are more dense than liquids.
- Ice is not more dense than water.

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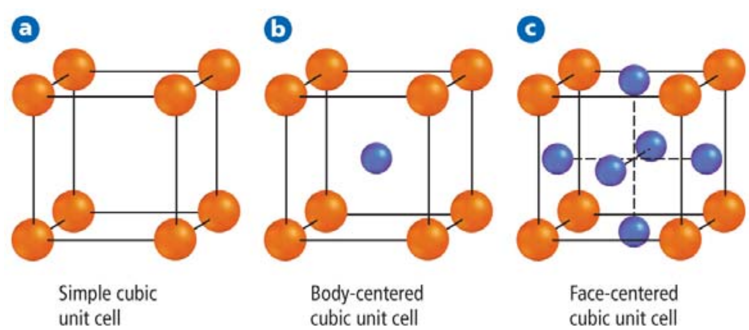
Solids (cont.)

- **Crystalline solids** are solids with atoms, ions, or molecules arranged in an orderly, geometric shape.



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Solids (cont.)

- *A* **unit cell** is the smallest arrangement of atoms in a crystal lattice that has the same symmetry as the whole crystal.

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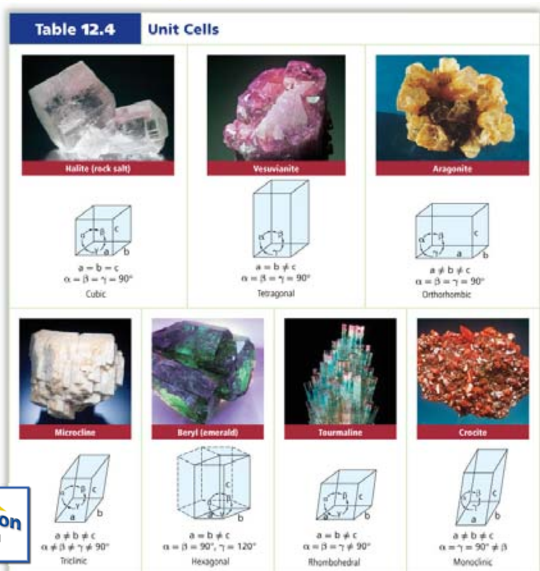
Solids (cont.)



Concepts In Motion
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Solids (cont.)

Table 12.5 Types of Crystalline Solids

Type	Unit Particles	Characteristics of Solid Phase	Examples
Atomic	atoms	soft to very soft; very low melting points; poor conductivity	group 18 elements
Molecular	molecules	fairly soft; low to moderately high melting points; poor conductivity	I_2, H_2O, NH_3, CO_2, $C_{12}H_{22}O_{11}$ (table sugar)
Covalent network	atoms connected by covalent bonds	very hard; very high melting points; often poor conductivity	diamond (C) and quartz (SiO_2)
Ionic	ions	hard; brittle; high melting points; poor conductivity	NaCl, KBr, $CaCO_3$
Metallic	atoms surrounded by mobile valence electrons	soft to hard; low to very high melting points; malleable and ductile; excellent conductivity	all metallic elements

Concepts In Motion
Click here to view an animated version of this graphic.

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Solids (cont.)

- **Amorphous solids** are solids in which the particles are not arranged in a regular, repeating pattern.
- Amorphous solids form when molten material cools quickly.

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Section 12.3 Assessment

CheckPoint

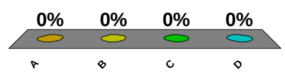
The smallest arrangement of atoms in a crystal that has the same pattern as the crystal is called ____.

A. crystal lattice

B. unit cell

C. crystalline

D. geometric cell



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CHAPTER 12

Section 12.3 Assessment

CheckPoint

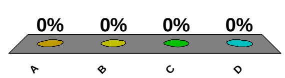
The viscosity of a liquid will increase as:

A. particle size decreases

B. temperature decreases

C. intermolecular forces decrease

D. particle size increases



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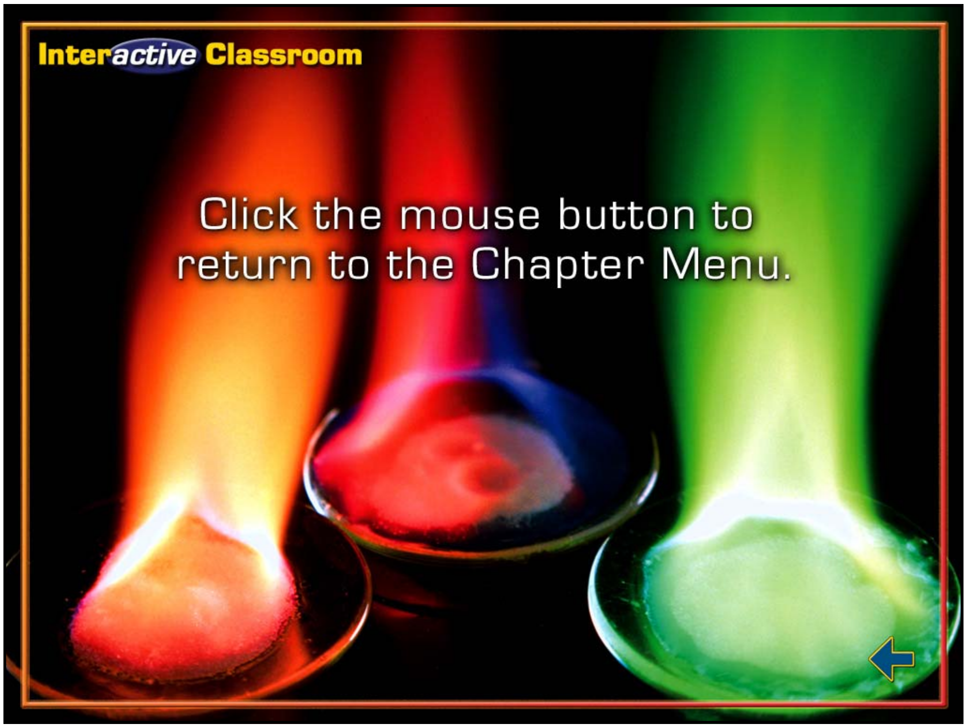

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Section 12.4 Phase Changes

Objectives

- **Explain** how the addition and removal of energy can cause a phase change.
- **Interpret** a phase diagram.

Review Vocabulary

phase change: a change from one state of matter to another

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Section 12.4 Phase Changes (cont.)

New Vocabulary

melting point
vaporization
evaporation
vapor pressure
boiling point
freezing point
condensation
deposition
phase diagram
triple point

MAIN Idea **Matter changes phase when energy is added or removed.**

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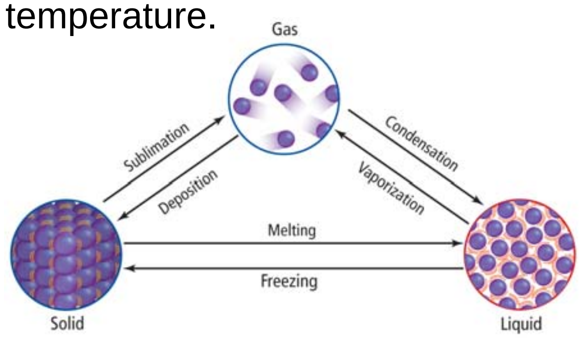
Phase Changes That Require Energy

- Melting occurs when heat flows into a solid object.
- Heat is the transfer of energy from an object at a higher temperature to an object at a lower temperature.



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Phase Changes That Require Energy (cont.)

- When ice is heated, the ice eventually absorbs enough energy to break the hydrogen bonds that hold the water molecules together.
- When the bonds break, the particles move apart and ice melts into water.
- The **melting point** of a crystalline solid is the temperature at which the forces holding the crystal lattice together are broken and it becomes a liquid.

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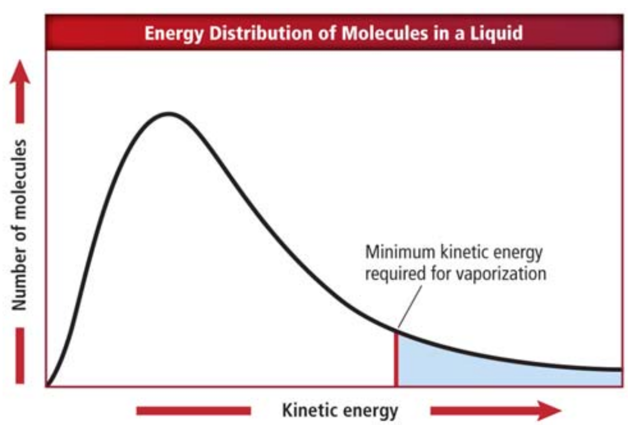
Phase Changes That Require Energy (cont.)

- Particles with enough energy escape from the liquid and enter the gas phase.



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Phase Changes That Require Energy (cont.)

- **Vaporization** is the process by which a liquid changes to a gas or vapor.
- **Evaporation** is vaporization only at the surface of a liquid.

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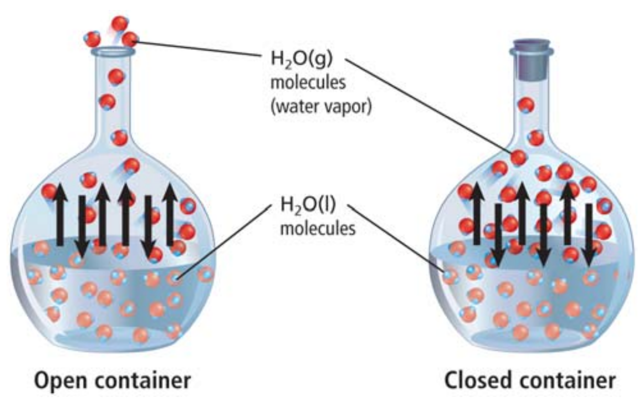
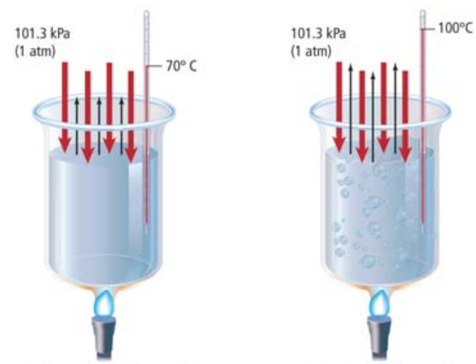
Phase Changes That Require Energy (cont.)

- In a closed container, the pressure exerted by a vapor over a liquid is called **vapor pressure**.



Phase Changes That Require Energy (cont.)

- The **boiling point** is the temperature at which the vapor pressure of a liquid equals the atmospheric pressure.



Below the boiling point At the boiling point

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Phase Changes That Require Energy (cont.)

- Sublimation is the process by which a solid changes into a gas without becoming a liquid.

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Phase Changes That Release Energy

- As heat flows from water to the surroundings, the particles lose energy.
- The **freezing point** is the temperature at which a liquid is converted into a crystalline solid.

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Phase Changes That Release Energy (cont.)

- As energy flows from water vapor, the velocity decreases.
- The process by which a gas or vapor becomes a liquid is called **condensation**.
- **Deposition** is the process by which a gas or vapor changes directly to a solid, and is the reverse of sublimation.

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Phase Diagrams

- A **phase diagram** is a graph of pressure versus temperature that shows in which phase a substance will exist under different conditions of temperature and pressure.

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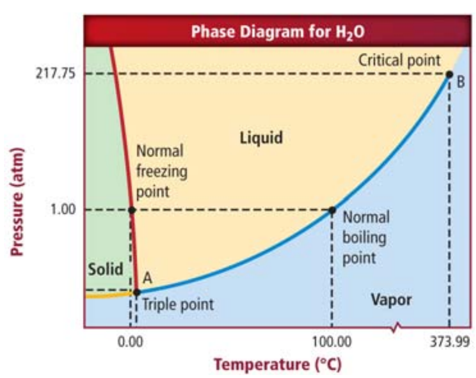
Phase Diagrams (cont.)

- The **triple point** is the point on a phase diagram that represents the temperature and pressure at which all three phases of a substance can coexist.



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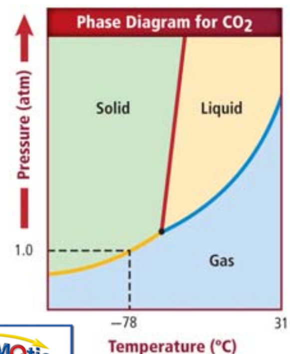
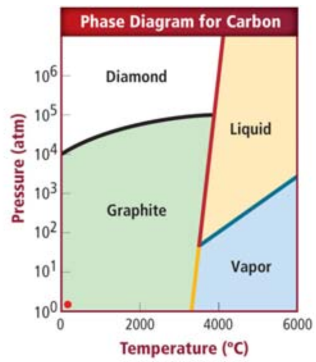
Phase Diagrams (cont.)

- The phase diagram for different substances are different from water.





Concepts In Motion
Click here to view an animated version of this graphic.

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Section 12.4 Assessment

CheckPoint

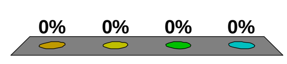
The addition of energy to water molecules will cause them to ____.

A. freeze

B. change to water vapor

C. form a crystal lattice

D. move closer together



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Section 12.4 Assessment

The transfer of energy from one object to another at a lower temperature is ____.

A. heat

B. degrees

C. conductivity

D. electricity

A	B	C	D
0%	0%	0%	0%

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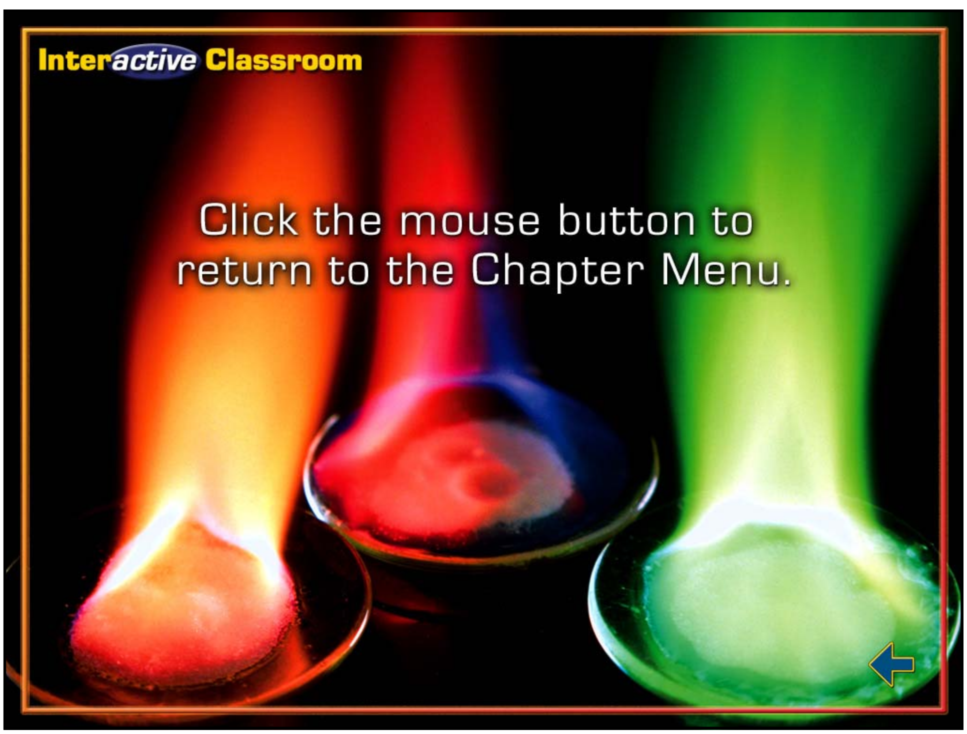

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