Preview

Lesson Starter

- Why did you not smell the odor of the vapor immediately?
- Explain this event in terms of the motion of molecules.

Objectives

- **State** the kinetic-molecular theory of matter, and **describe** how it explains certain properties of matter.
- **List** the five assumptions of the kinetic-molecular theory of gases. **Define** the terms ideal gas and real gas.
- **Describe** each of the following characteristic properties of gases: expansion, density, fluidity, compressibility, diffusion, and effusion.
- **Describe** the conditions under which a real gas deviates from "ideal" behavior.

- The **kinetic-molecular theory** is based on the idea that particles of matter are always in motion.
- The theory can be used to explain the properties of solids, liquids, and gases in terms of the energy of particles and the forces that act between them.

End Of Slide

The Kinetic-Molecular Theory of Gases

- An **ideal gas** is a hypothetical gas that perfectly fits all the assumptions of the kinetic-molecular theory.

- The kinetic-molecular theory of gases is based on the following five assumptions:

  1. Gases consist of large numbers of tiny particles that are far apart relative to their size.

     - Most of the volume occupied by a gas is empty space

The Kinetic-Molecular Theory of Gases, *continued*

2. Collisions between gas particles and between particles and container walls are elastic collisions.

   - An **elastic collision** is one in which there is no net loss of total kinetic energy.

3. Gas particles are in continuous, rapid, random motion. They therefore possess kinetic energy, which is energy of motion.

The Kinetic-Molecular Theory of Gases, *continued*

4. There are no forces of attraction between gas particles.

5. The temperature of a gas depends on the average kinetic energy of the particles of the gas.

- The kinetic energy of any moving object is given by the following equation:

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

End Of Slide

The Kinetic-Molecular Theory of Gases, *continued*

- All gases at the same temperature have the same average kinetic energy.
- At the same temperature, lighter gas particles, have higher average speeds than do heavier gas particles.
  - Hydrogen molecules will have a higher speed than oxygen molecules.

Kinetic Molecular Theory

Click below to watch the Visual Concept.

Visual Concept

Properties of Gases

Click below to watch the Visual Concept.

Visual Concept

The Kinetic-Molecular Theory and the Nature of Gases

- The kinetic-molecular theory applies only to ideal gases.
- Many gases behave nearly ideally if pressure is not very high and temperature is not very low.

Expansion

- Gases do not have a definite shape or a definite volume.
  - They completely fill any container in which they are enclosed.

The Kinetic-Molecular Theory and the Nature of Gases, *continued*

Expansion, *continued*

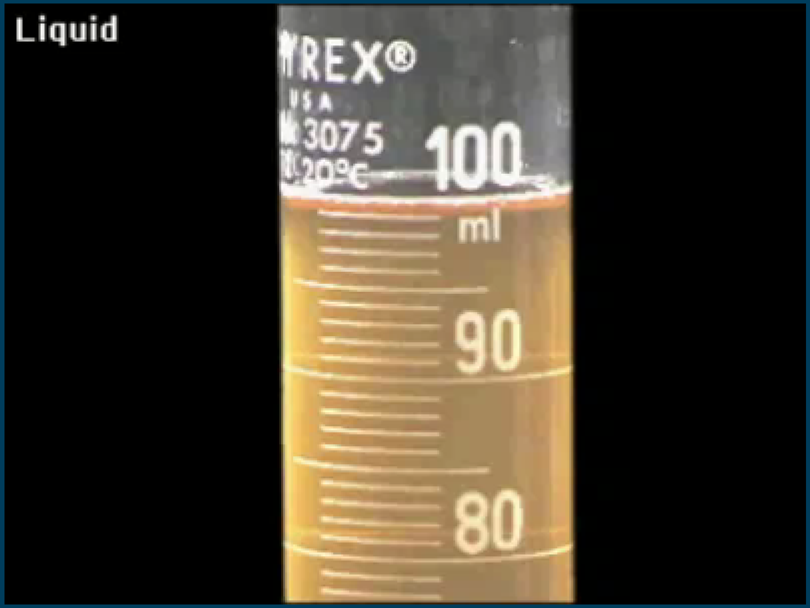
- Gas particles move rapidly in all directions (assumption 3) without significant attraction between them (assumption 4).

Fluidity

- Because the attractive forces between gas particles are insignificant (assumption 4), gas particles glide easily past one another.
  - Because liquids and gases flow, they are both referred to as *fluids.*

Fluid

Click below to watch the Visual Concept.



The Kinetic-Molecular Theory and the Nature of Gases, *continued*

Low Density

- The density of a gaseous substance at atmospheric pressure is about 1/1000 the density of the same substance in the liquid or solid state.
  - The reason is that the particles are so much farther apart in the gaseous state (assumption 1).

Compressibility

- During compression, the gas particles, which are initially very far apart (assumption 1), are crowded closer together.

End Of Slide

The Kinetic-Molecular Theory and the Nature of Gases, *continued*

Diffusion and Effusion

- Gases spread out and mix with one another, even without being stirred.
  - The random and continuous motion of the gas molecules (assumption 3) carries them throughout the available space.
- Such spontaneous mixing of the particles of two substances caused by their random motion is called **diffusion.**

End of Slide

The Kinetic-Molecular Theory and the Nature of Gases, *continued*

Diffusion and Effusion, *continued*

- **Effusion** is a process by which gas particles pass through a tiny opening.

- The rates of effusion of different gases are directly proportional to the velocities of their particles.

  - Molecules of low mass effuse faster than molecules of high mass.

End Of Slide

Comparing Diffusion and Effusion

Click below to watch the Visual Concept.

Visual Concept

Deviations of Real Gases from Ideal Behavior

- Because particles of gases occupy space and exert attractive forces on each other, all real gases deviate to some degree from ideal gas behavior.
- A **real gas** is a gas that does not behave completely according to the assumptions of the kinetic-molecular theory.
- At very high pressures and low temperatures, a gas is most likely to behave like a non-ideal gas.
- The more polar a gas's molecules are, the more the gas will deviate from ideal gas behavior.

Preview

- Lesson Starter
- Objectives
- Properties of Liquids and the Kinetic-Molecular Theory

Lesson Starter

- How are you able to tell that the container is filled with a liquid?
- Liquids have definite volume but take the shape of their container.
- How is this different from gases?
- Gases do not have a fixed shape or a fixed volume.

End of Slide

Objectives

- **Describe** the motion of particles in liquids and the properties of liquids according to the kinetic-molecular theory.
- **Discuss** the process by which liquids can change into a gas. **Define** *vaporization.*
- **Discuss** the process by which liquids can change into a solid. **Define** *freezing.*

Properties of Liquids and the Kinetic-Molecular Theory

- A liquid can be described as a form of matter that has a definite volume and takes the shape of its container.
- The attractive forces between particles in a liquid are more effective than those between particles in a gas.
- This attraction between liquid particles is caused by the intermolecular forces:
  - dipole-dipole forces
  - London dispersion forces
  - hydrogen bonding

Properties of Liquids and the Kinetic-Molecular Theory, *continued*

- The particles in a liquid are not bound together in fixed positions. Instead, they move about constantly.

- A **fluid** is a substance that can flow and therefore take the shape of its container.

Relatively High Density

- At normal atmospheric pressure, most substances are hundreds of times denser in a liquid state than in a gaseous state.

End Of Slide

Properties of Liquids and the Kinetic-Molecular Theory, *continued*

Relative Incompressibility

- Liquids are much less compressible than gases because liquid particles are more closely packed together.

Ability to Diffuse

- Any liquid gradually diffuses throughout any other liquid in which it can dissolve.
    - The constant, random motion of particles causes diffusion in liquids.

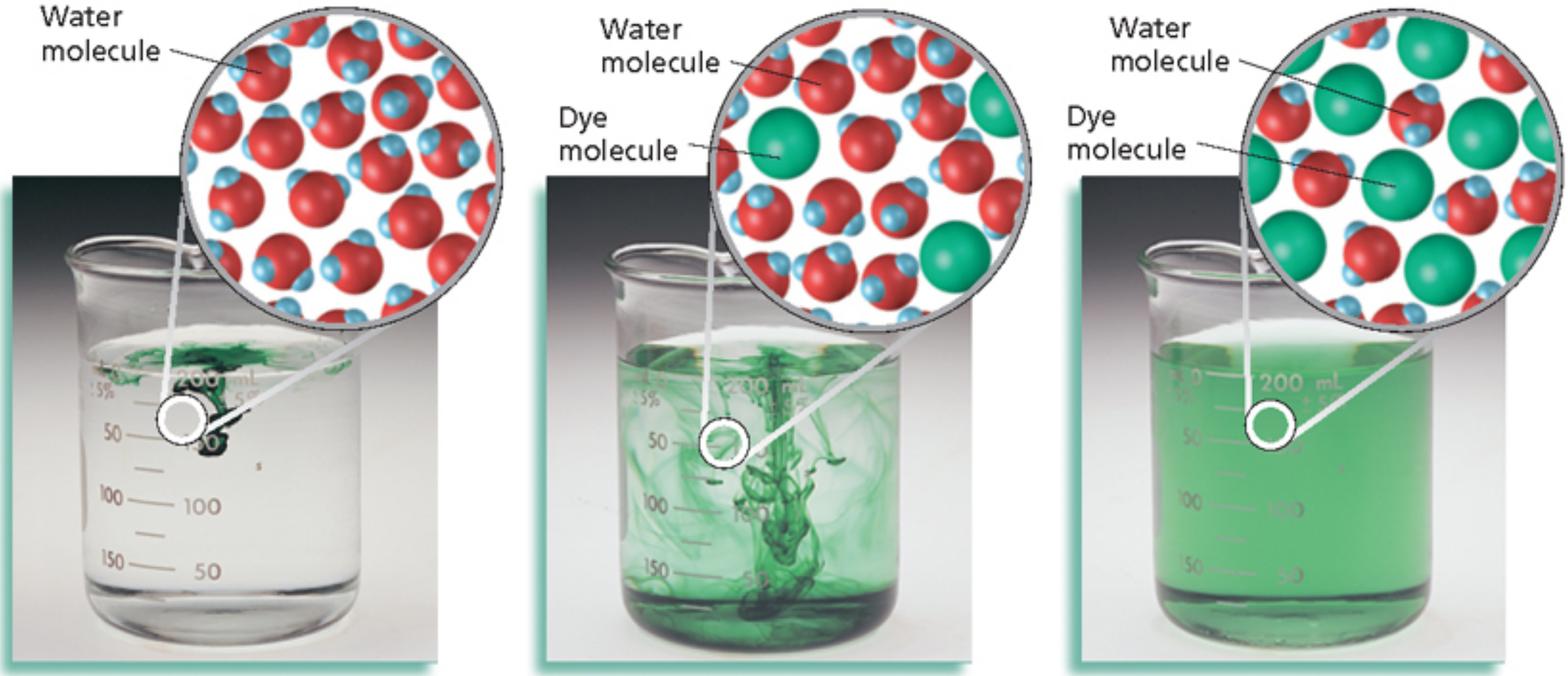
Properties of Liquids and the Kinetic-Molecular Theory, *continued*

Ability to Diffuse

- Diffusion is much slower in liquids than in gases.
  - Liquid particles are closer together.
  - The attractive forces between the particles of a liquid slow their movement.
  - As the temperature of a liquid is increased, diffusion occurs more rapidly.

End Of Slide

Diffusion



Like gases, the two liquids in this beaker diffuse over time. The green liquid food coloring from the drop will eventually form a uniform solution with the water.

Diffusion in a Liquid

Click below to watch the Visual Concept.

Visual Concept

Properties of Liquids and the Kinetic-Molecular Theory, *continued*

Surface Tension

- A property common to all liquids is **surface tension,** a force that tends to pull adjacent parts of a liquid's surface together, thereby decreasing surface area to the smallest possible size.
- The higher the force of attraction between the particles of a liquid, the higher the surface tension.
- The molecules at the surface of the water can form hydrogen bonds with the other water, but not with the molecules in the air above them.

Surface Tension

Click below to watch the Visual Concept.

Visual Concept

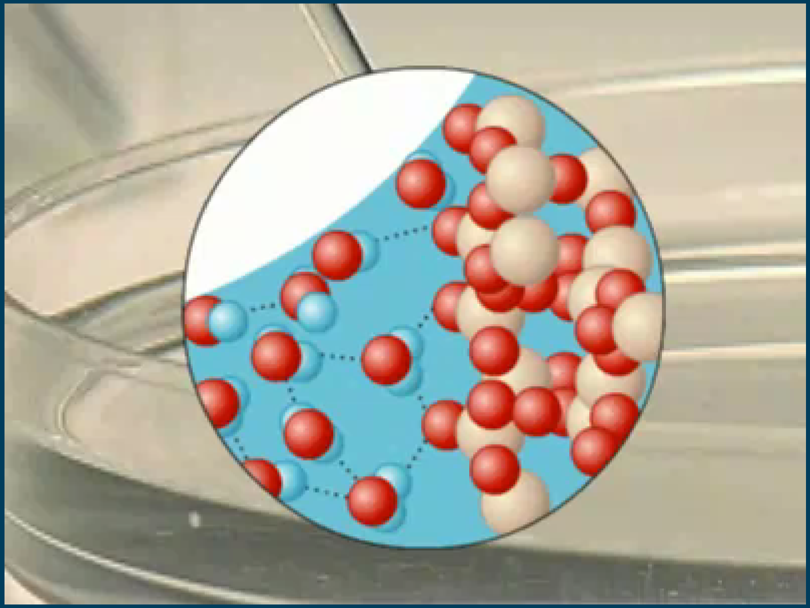
Properties of Liquids and the Kinetic-Molecular Theory, *continued*

Surface Tension, *continued*

- **Capillary action** is the attraction of the surface of a liquid to the surface of a solid.
- This attraction tends to pull the liquid molecules upward along the surface and against the pull of gravity.
- The same process is responsible for the concave liquid surface, called a *meniscus*, that forms in a test tube or graduated cylinder.

Capillary Action

Click below to watch the Visual Concept.



Properties of Liquids and the Kinetic-Molecular Theory, *continued*

Evaporation and Boiling

- The process by which a liquid or solid changes to a gas is **vaporization.**
- **Evaporation** is the process by which particles escape from the surface of a nonboiling liquid and enter the gas state.
    - *Boiling* is the change of a liquid to bubbles of vapor that appear throughout the liquid.
- Evaporation occurs because the particles of a liquid have different kinetic energies.

Properties of Liquids and the Kinetic-Molecular Theory, *continued*

Formation of Solids

- When a liquid is cooled, the average energy of its particles decreases.

- The physical change of a liquid to a solid by removal of energy as heat is called **freezing** or solidification.

Freezing

Click below to watch the Visual Concept.

Visual Concept

Preview

- Lesson Starter
- Objectives
- Properties of Solids and the Kinetic-Molecular Theory
- Crystalline Solids

Lesson Starter

- Compare the plaster of Paris mixture before it hardens to the product after it hardens.

End of Slide

< Back Next > Preview Main

Objectives

- **Describe** the motion of particles in solids and the properties of solids according to the kinetic-molecular theory.
- **Distinguish** between the two types of solids.
- **Describe** the different types of crystal symmetry.
- **Define** *crystal structure* and *unit cell.*

End Of Slide

Properties of Solids and the Kinetic-Molecular Theory

- The particles of a solid are more closely packed than those of a liquid or gas.
- All interparticle attractions exert stronger effects in solids than in the corresponding liquids or gases.
- Attractive forces tend to hold the particles of a solid in relatively fixed positions.
  - Solids are more ordered than liquids and are much more ordered than gases.

Properties of Solids and the Kinetic-Molecular Theory, *continued*

- There are two types of solids: crystalline solids and amorphous solids.
- Most solids are **crystalline solids**—*they consist of crystals.*
  - A **crystal** is a substance in which the particles are arranged in an orderly, geometric, repeating pattern.
- An **amorphous solid** is one in which the particles are arranged randomly.

Properties of Solids and the Kinetic-Molecular Theory, *continued*

Definite Shape and Volume

- Solids can maintain a definite shape without a container.
- Crystalline solids are geometrically regular.
- The volume of a solid changes only slightly with a change in temperature or pressure.
  - Solids have definite volume because their particles are packed closely together.

Properties of Solids and the Kinetic-Molecular Theory, *continued*

Definite Melting Point

- **Melting** is the physical change of a solid to a liquid by the addition of energy as heat.
- The temperature at which a solid becomes a liquid is its **melting point.**
  - At this temperature, the kinetic energies of the particles within the solid overcome the attractive forces holding them together.

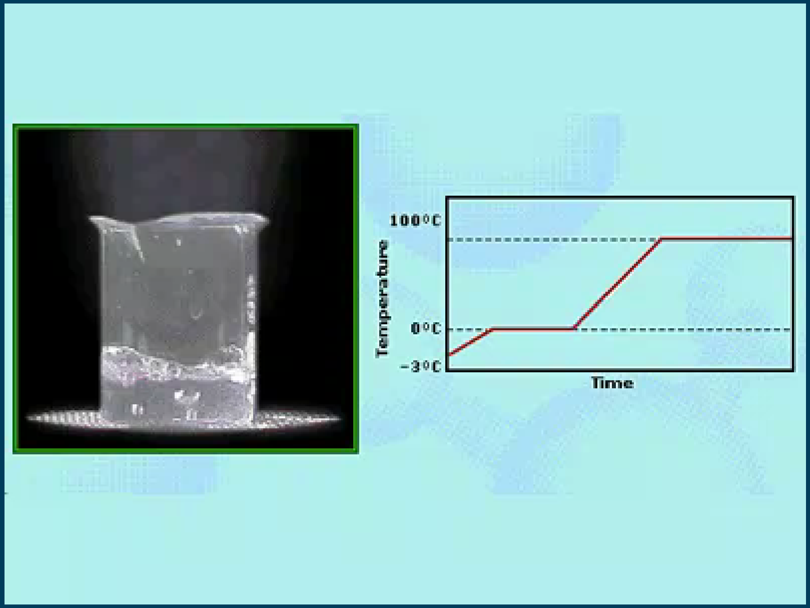
End Of Slide

Melting Point

Click below to watch the Visual Concept.

Melting

Click below to watch the Visual Concept.



Properties of Solids and the Kinetic-Molecular Theory, *continued*

Definite Melting Point, *continued*

- Amorphous solids have no definite melting point.
  - example: glass and plastics
- Amorphous solids are sometimes classified as **supercooled liquids,** which are substances that retain certain liquid properties even at temperatures at which they appear to be solid.
  - These properties exist because the particles in amorphous solids are arranged randomly.

Properties of Solids and the Kinetic-Molecular Theory, *continued*

High Density and Incompressibility

- In general, substances are most dense in the solid state.
  - The higher density results from the fact that the particles of a solid are more closely packed than those of a liquid or a gas.
- For practical purposes, solids can be considered incompressible.

Properties of Solids and the Kinetic-Molecular Theory, *continued*

Low Rate of Diffusion

- The rate of diffusion is millions of times slower in solids than in liquids

Crystalline Solids

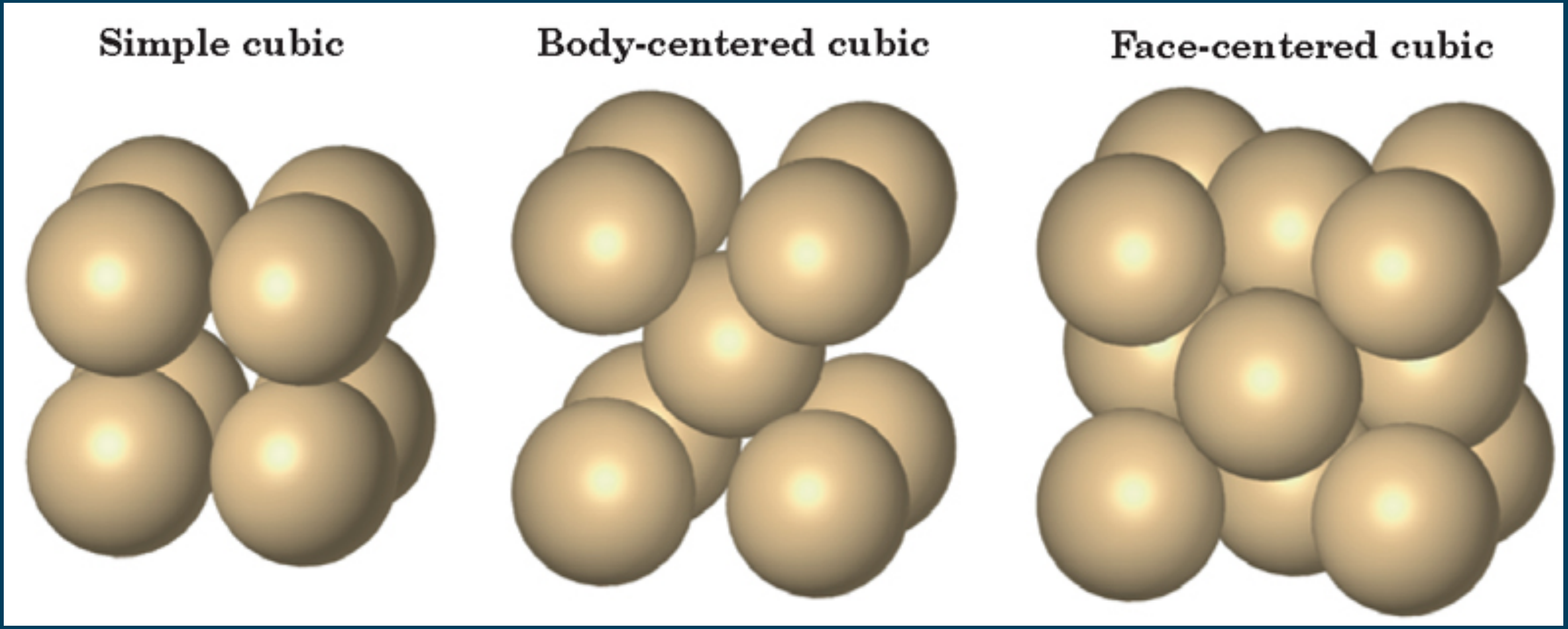
- Crystalline solids exist either as single crystals or as groups of crystals fused together.

- The total three-dimensional arrangement of particles of a crystal is called a **crystal structure.**

  - The arrangement of particles in the crystal can be represented by a coordinate system called a *lattice.*

- The smallest portion of a crystal lattice that shows the three-dimensional pattern of the entire lattice is called a **unit cell.**

Unit Cells



Types of Basic Crystalline Systems

Click below to watch the Visual Concept.

Visual Concept

Types of Crystals

Click below to watch the Visual Concept.

Visual Concept

Crystalline Solids, *continued*

- A crystal and its unit cells can have any one of seven types of symmetry.

Binding Forces in Crystals

- Crystal structures can also be described in terms of the types of particles in them and the types of chemical bonding between the particles.

Crystalline Solids, *continued*

Binding Forces in Crystals, *continued*

- Melting and Boiling Points of Representative Crystaline Solids

Type of substance	Formula	Melting point (°C)	Boiling point at 1 atm (°C)
Ionic	NaCl	801	1413
	MgF_2	1266	2239
Covalent network	$(SiO_2)_x$	1610	2230
	C_x (diamond)	3500	3930
Metallic	Hg	–39	357
	Cu	1083	2567
	Fe	1535	2750
	W	3410	5660
Covalent molecular (nonpolar)	H_2	–259	–253
	O_2	–218	–183
	CH_4	–182	–164
	CCl_4	–23	77
	C_6H_6	6	80
Covalent molecular (polar)	NH_3	–78	–33
	H_2O	0	100

Crystalline Solids, *continued*

Binding Forces in Crystals, *continued*

1. *Ionic crystals*—The ionic crystal structure consists of positive and negative ions arranged in a regular pattern.

   - Generally, ionic crystals form when Group 1 or Group 2 metals combine with Group 16 or Group 17 nonmetals or nonmetallic polyatomic ions.
   - These crystals are hard and brittle, have high melting points, and are good insulators.

Crystalline Solids, *continued*

Binding Forces in Crystals, *continued*

2. Covalent network crystals—In covalent network crystals, each atom is covalently bonded to its nearest neighboring atoms.

- The covalent bonding extends throughout a network that includes a very large number of atoms.

- The network solids are very hard and brittle, have high melting points and are usually nonconductors or semiconductors.

End Of Slide

Crystalline Solids, *continued*

Binding Forces in Crystals, *continued*

3. Metallic crystals—The metallic crystal structure consists of metal cations surrounded by a sea of delocalized valence electrons.

- The electrons come from the metal atoms and belong to the crystal as a whole.

- The freedom of these delocalized electrons to move throughout the crystal explains the high electric conductivity of metals.

End Of Slide

Crystalline Solids, *continued*

Binding Forces in Crystals, *continued*

4. *Covalent molecular crystals*—The crystal structure of a covalent molecular substance consists of covalently bonded molecules held together by intermolecular forces.

- If the molecules are nonpolar, then there are only weak London dispersion forces between molecules.
- In a polar covalent molecular crystal, molecules are held together by dispersion forces, by dipole-dipole forces, and sometimes by hydrogen bonding.

Crystalline Solids, *continued*

Binding Forces in Crystals, *continued*

4. Covalent molecular crystals, continued

- Covalent molecular crystals have low melting points, are easily vaporized, are relatively soft, and are good insulators.

Amorphous Solids

- The word *amorphous* comes from the Greek for "without shape."
- Unlike the atoms that form crystals, the atoms that make up amorphous solids are not arranged in a regular pattern.

Comparing Cohesion and Adhesion

Click below to watch the Visual Concept.

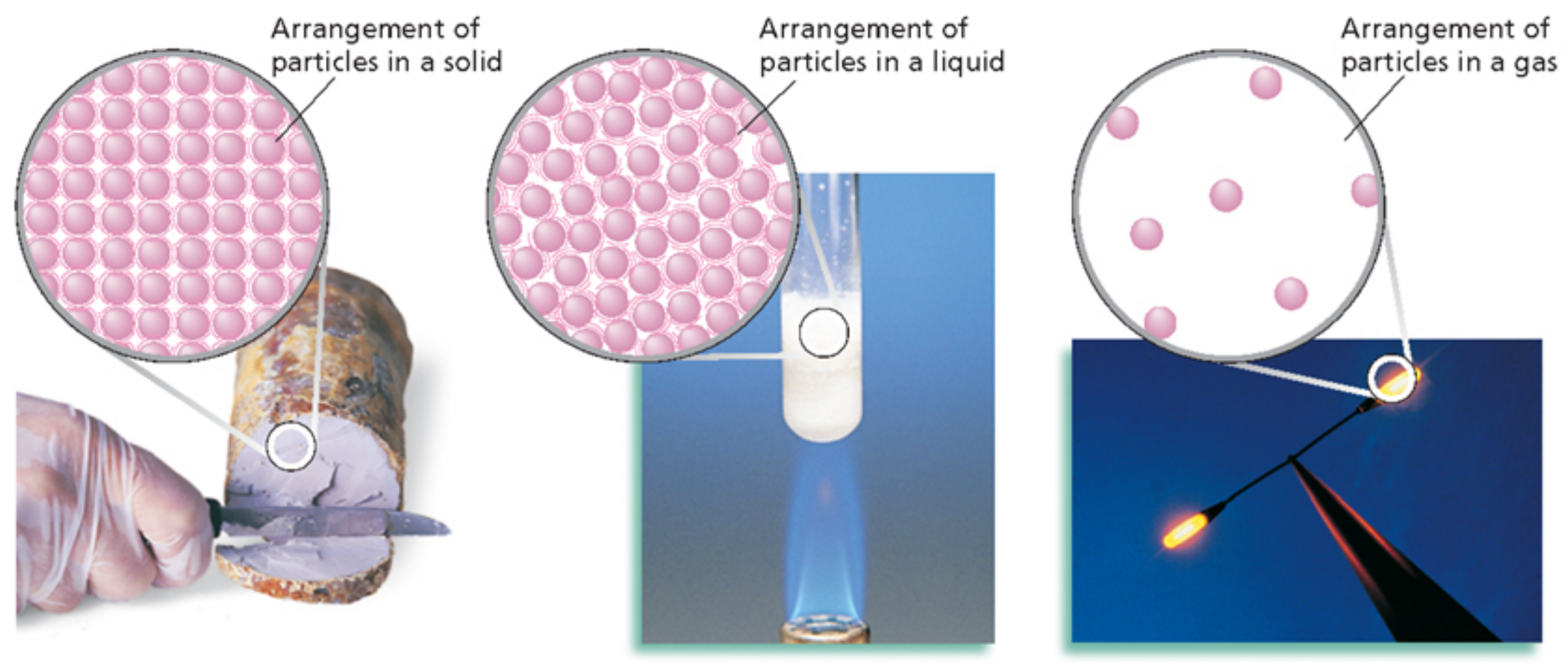
Visual Concept

Vaporization and Condensation

Click below to watch the Visual Concept.

Visual Concept

Sodium as a Solid, Liquid, and Gas



Preview

- Lesson Starter
- Objectives
- Changes of State and Equilibrium
- Equilibrium Vapor Pressure of a Liquid
- Boiling
- Freezing and Melting
- Phase Diagrams

Lesson Starter

- Why does the balloon inflate after the solid dry ice is added?
- The solid CO_2 sublimes to form CO_2 gas.
- The gas occupies more volume than the solid.

Objectives

- **Explain** the relationship between equilibrium and changes of state.

- **Interpret** phase diagrams.

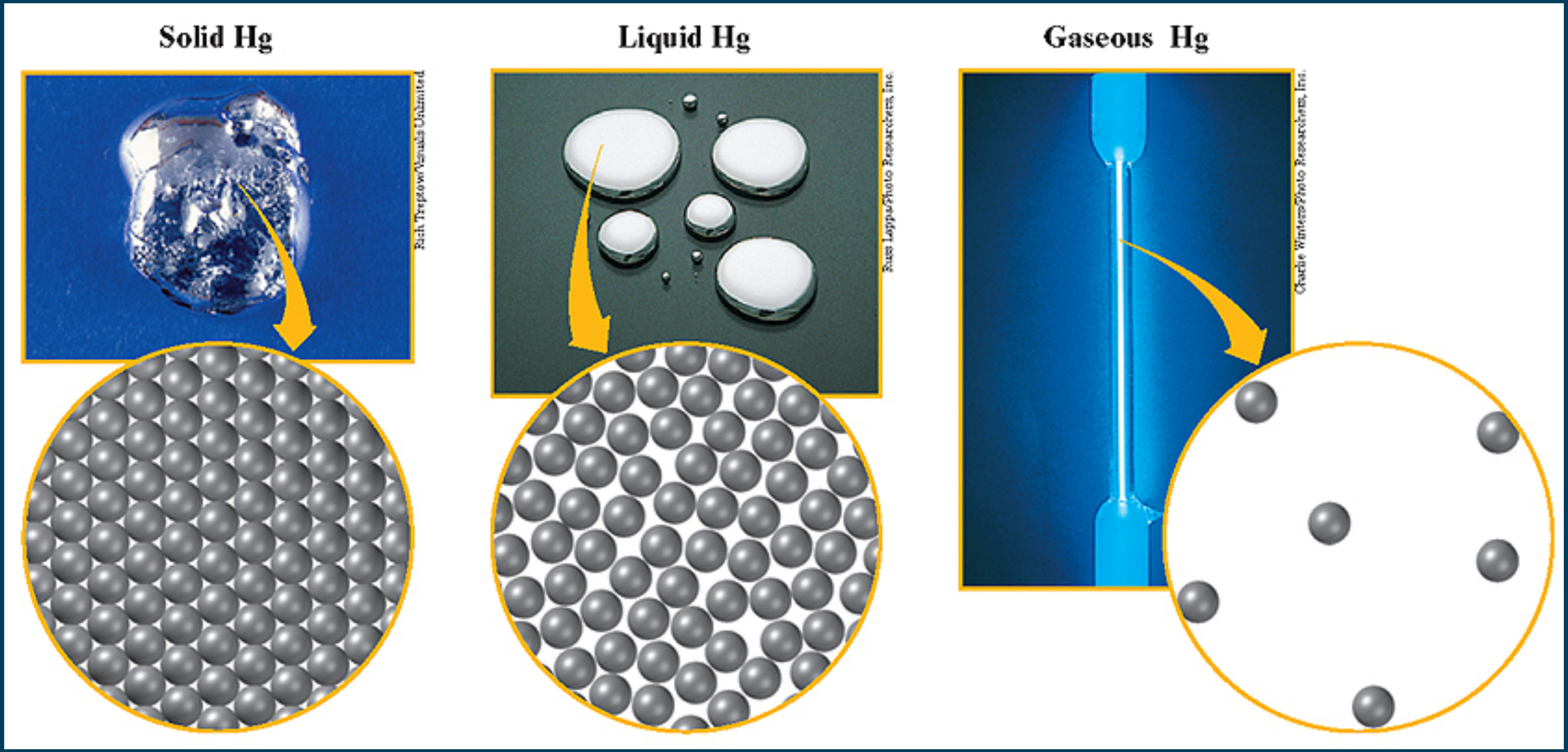
- **Explain** what is meant by equilibrium vapor pressure.

- **Describe** the processes of boiling, freezing, melting, and sublimation.

Possible Changes of State

Change of state	Process	Example
Solid ⟶ liquid	melting	ice ⟶ water
Solid ⟶ gas	sublimation	dry ice ⟶ CO_2 gas
Liquid ⟶ solid	freezing	water ⟶ ice
Liquid ⟶ gas	vaporization	liquid bromine ⟶ bromine vapor
Gas ⟶ liquid	condensation	water vapor ⟶ water
Gas ⟶ solid	deposition	water vapor ⟶ ice

Mercury in Three States



Changes of State and Equilibrium

- A **phase** is any part of a system that has uniform composition and properties.
- **Condensation** is the process by which a gas changes to a liquid.
- A gas in contact with its liquid or solid phase is often called a *vapor.*
- **Equilibrium** is a dynamic condition in which two opposing changes occur at equal rates in a closed system.

Equilibrium

Click below to watch the Visual Concept.

Visual Concept

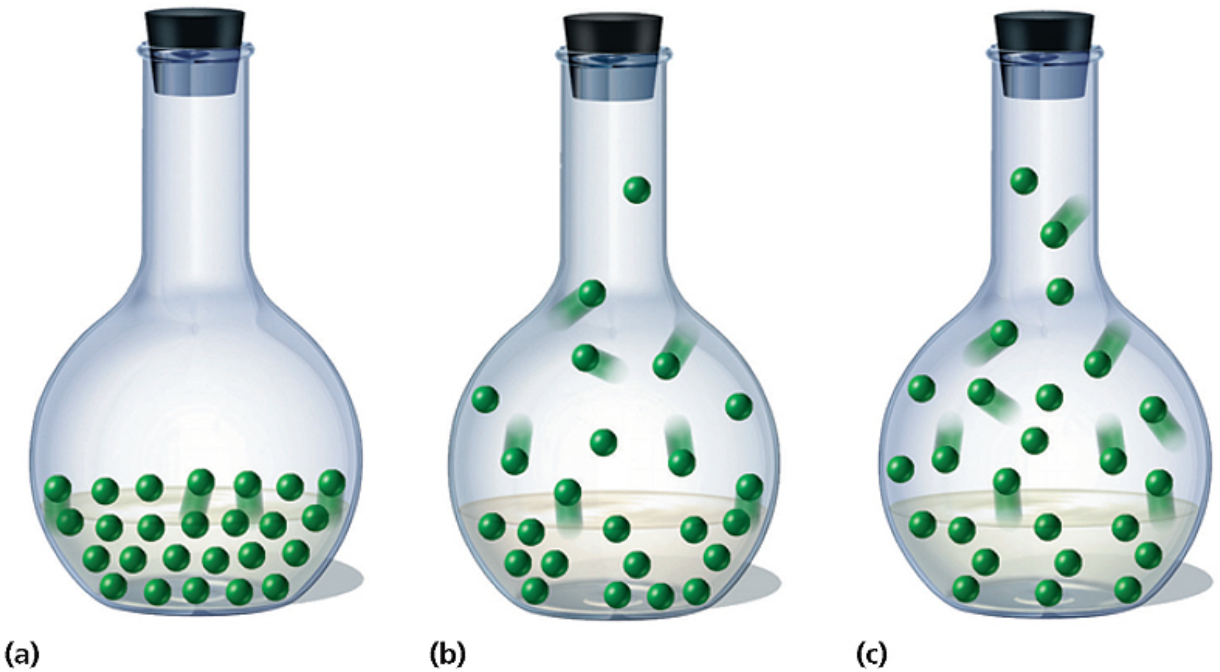
Changes of State and Equilibrium, *continued*

- Eventually, in a closed system, the rate of condensation equals the rate of evaporation, and a state of equilibrium is established.

End Of Slide

< Back Next > Preview Main

Liquid-Vapor Equilibrium System



A liquid-vapor equilibrium develops in a closed system. (a) At first there is only liquid present, but molecules are beginning to evaporate. (b) Evaporation continues at a constant rate. Some vapor molecules are beginning to condense to liquid. (c) Equilibrium has been reached between the rate of condensation and the rate of evaporation.

Equilibrium Vapor Pressure of a Liquid

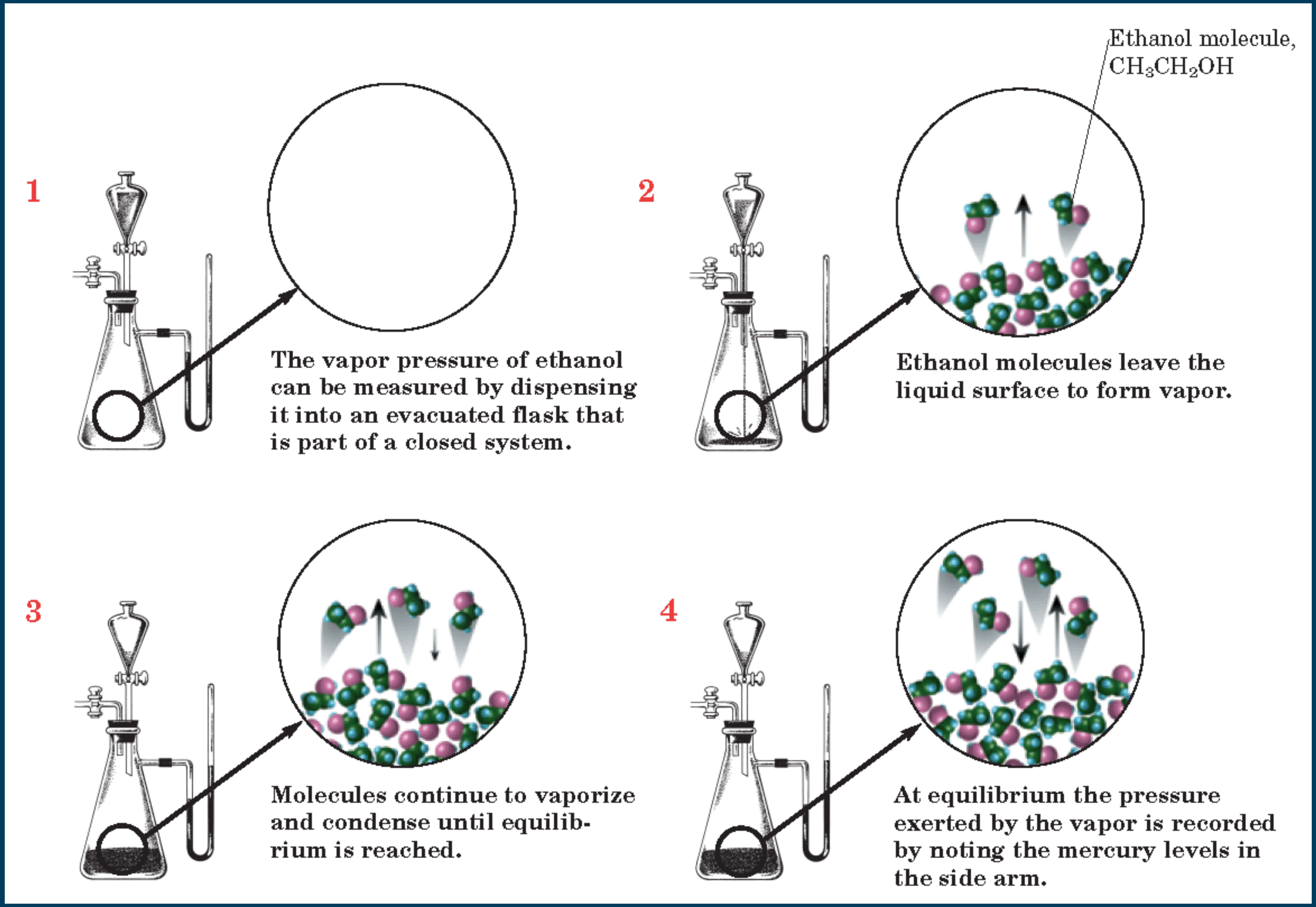
- Vapor molecules in equilibrium with a liquid in a closed system exert a pressure proportional to the concentration of molecules in the vapor phase.
- The pressure exerted by a vapor in equilibrium with its corresponding liquid at a given temperature is called the **equilibrium vapor pressure** of the liquid.
- The equilibrium vapor pressure increases with increasing temperature.
  - Increasing the temperature of a liquid increases the average kinetic energy of the liquid's molecules.

Equilibrium and Vapor Pressure

Click below to watch the Visual Concept.

Visual Concept

Measuring the Vapor Pressure of a Liquid



Factors Affecting Equilibrium

Click below to watch the Visual Concept.

Visual Concept

Shifts in the Equilibrium Due to the Application of Heat

Click below to watch the Visual Concept.

Visual Concept

Equilibrium Vapor Pressure of a Liquid, *continued*

- Every liquid has a specific equilibrium vapor pressure at a given temperature.
  - All liquids have characteristic forces of attraction between their particles.
- **Volatile liquids** are liquids that evaporate readily.
  - They have relatively weak forces of attraction between their particles.
  - example: ether

End of Slide

Equilibrium Vapor Pressure of a Liquid, *continued*

- Nonvolatile liquids do not evaporate readily.
  - They have relatively strong attractive forces between their particles.
  - example: molten ionic compounds

Comparing Volatile and Nonvolatile Liquids

Click below to watch the Visual Concept.

Visual Concept

Boiling

- **Boiling** is the conversion of a liquid to a vapor within the liquid as well as at its surface.
- The **boiling point** of a liquid is the temperature at which the equilibrium vapor pressure of the liquid equals the atmospheric pressure.
  - The lower the atmospheric pressure is, the lower the boiling point is.

Boiling, *continued*

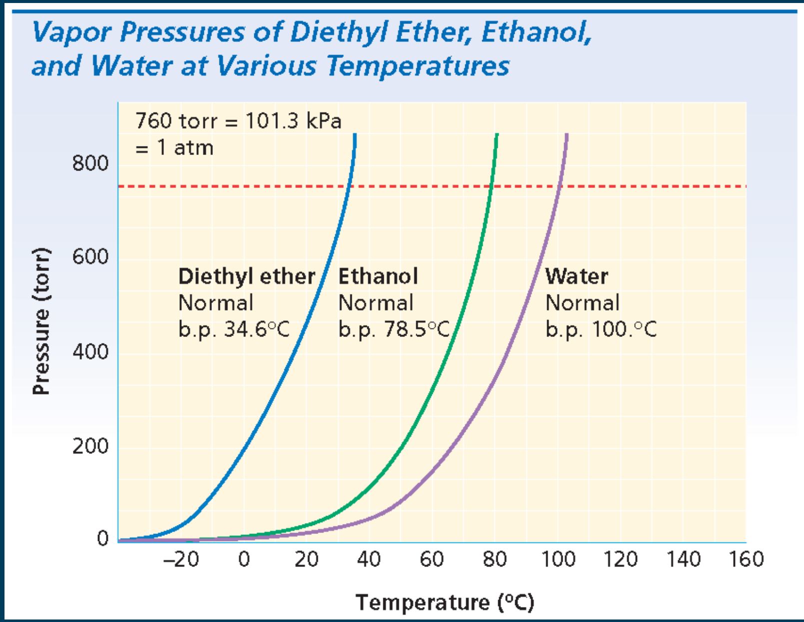
- At the boiling point, all of the energy absorbed is used to evaporate the liquid, and the temperature remains constant as long as the pressure does not change.

- If the pressure above the liquid being heated is increased, the temperature of the liquid will rise until the vapor pressure equals the new pressure and the liquid boils once again.

End Of Slide

Boiling, *continued*

- The *normal* boiling point of a liquid is the boiling point at normal atmospheric pressure (1 atm, 760 torr, or 101.3 kPa).
  - The normal boiling point of water is exactly 100°C.



Boiling, *continued*

Energy and Boiling

- Energy must be added continuously in order to keep a liquid boiling
- The temperature at the boiling point remains constant despite the continuous addition of energy.
    - The added energy is used to overcome the attractive forces between molecules of the liquid during the liquid-to-gas change and is stored in the vapor as potential energy.

Boiling, *continued*

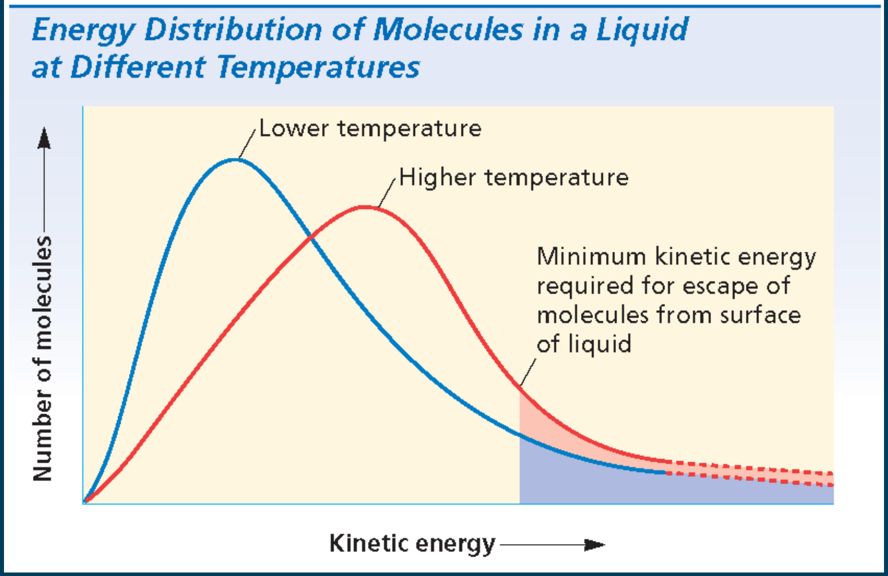
Molar Enthalpy of Vaporization

- The amount of energy as heat that is needed to vaporize one mole of liquid at the liquid's boiling point at constant pressure is called the liquid's **molar enthalpy of vaporization,** ΔH_v.
- The magnitude of the molar enthalpy of vaporization is a measure of the attraction between particles of the liquid.
  - The stronger this attraction is, the higher molar enthalpy of vaporization.

Boiling, *continued*

Molar Enthalpy of Vaporization, *continued*

- Each liquid has a characteristic molar enthalpy of vaporization.
  - Water has an unusually high molar enthalpy of vaporization due to hydrogen bonding in liquid water.



Freezing and Melting

- The physical change of a liquid to a solid is called **freezing.**
- Freezing involves a loss of energy in the form of heat by the liquid.

$$\text{liquid} \longrightarrow \text{solid} + \text{energy}$$

- In the case of a pure crystalline substance, this change occurs at constant temperature.

End Of Slide

Freezing and Melting, *continued*

- The normal **freezing point** is the temperature at which the solid and liquid are in equilibrium at 1 atm (760 torr, or 101.3 kPa) pressure.
  - At the freezing point, particles of the liquid and the solid have the same average kinetic energy.
- Melting, the reverse of freezing, also occurs at constant temperature.

$$\text{solid + energy} \longrightarrow \text{liquid}$$

End of Slide

Freezing and Melting, *continued*

- At equilibrium, melting and freezing proceed at equal rates.

$$\text{solid + energy} \rightleftharpoons \text{liquid}$$

- At normal atmospheric pressure, the temperature of a system containing ice and liquid water will remain at 0.°C as long as both ice and water are present.
    - Only after all the ice has melted will the addition of energy increase the temperature of the system.

End Of Slide

Freezing and Melting, *continued*

Molar Enthalpy of Fusion

- The amount of energy as heat required to melt one mole of solid at the solid's melting point is the solid's **molar enthalpy of fusion,** ΔH_f.

- The magnitude of the molar enthalpy of fusion depends on the attraction between the solid particles.

Freezing and Melting, *continued*

Sublimation and Deposition

- At sufficiently low temperature and pressure conditions, a liquid cannot exist.
  - Under such conditions, a solid substance exists in equilibrium with its vapor instead of its liquid.

$$\text{solid} + \text{energy} \rightleftharpoons \text{vapor}$$

- The change of state from a solid directly to a gas is known as **sublimation.**
- The reverse process is called **deposition,** the change of state from a gas directly to a solid.

End Of Slide

Comparing Sublimation and Deposition

Click below to watch the Visual Concept.

Visual Concept

Phase Diagrams

- A **phase diagram** is a graph of pressure versus temperature that shows the conditions under which the phases of a substance exist.
- The **triple point** of a substance indicates the temperature and pressure conditions at which the solid, liquid, and vapor of the substance can coexist at equilibrium.
- The **critical point** of a substance indicates the critical temperature and critical pressure.

Phase Diagrams

- The **critical temperature** (t_c) is the temperature above which the substance cannot exist in the liquid state.
  - Above this temperature, water cannot be liquefied, no matter how much pressure is applied.
- The **critical pressure** (P_c) is the lowest pressure at which the substance can exist as a liquid at the critical temperature.

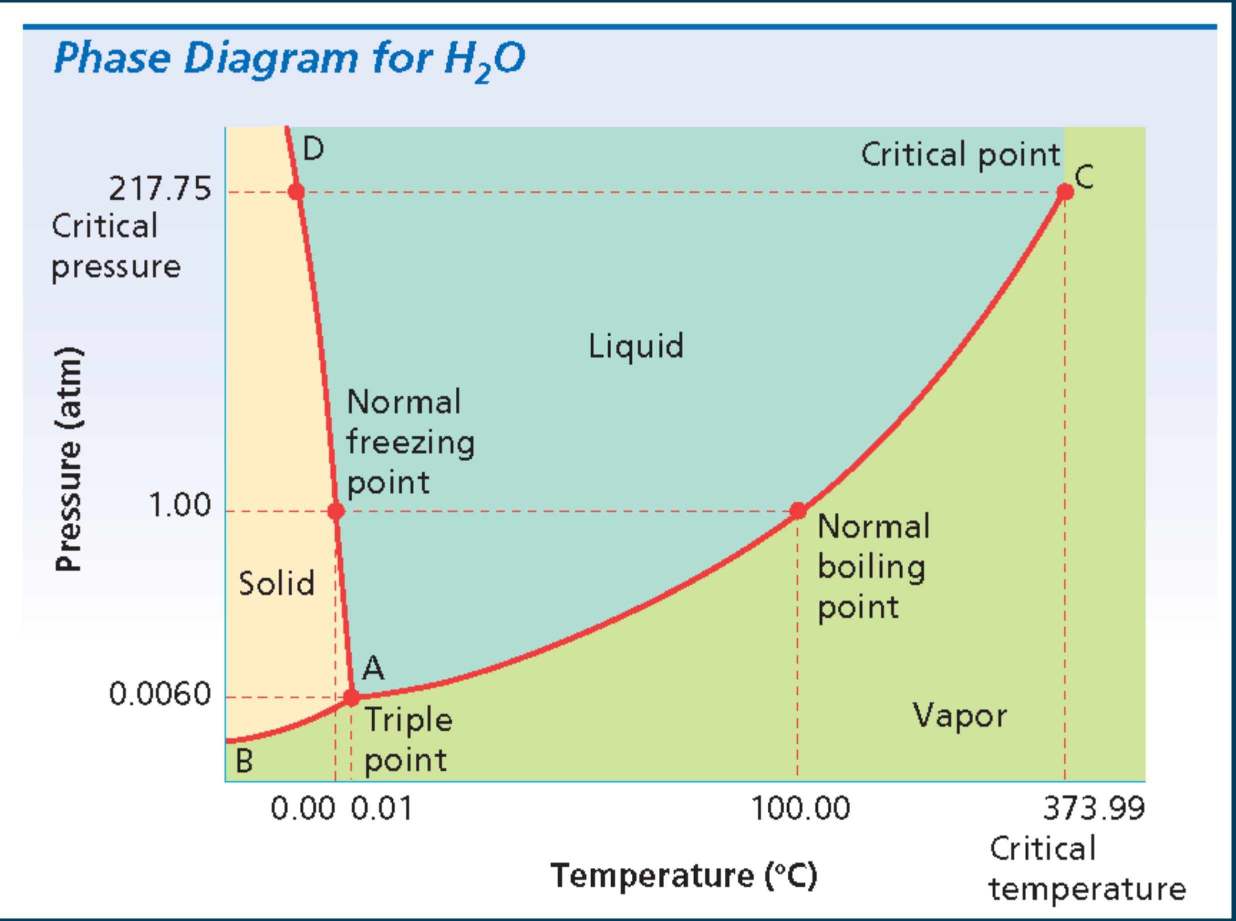
End Of Slide

Phase Diagram

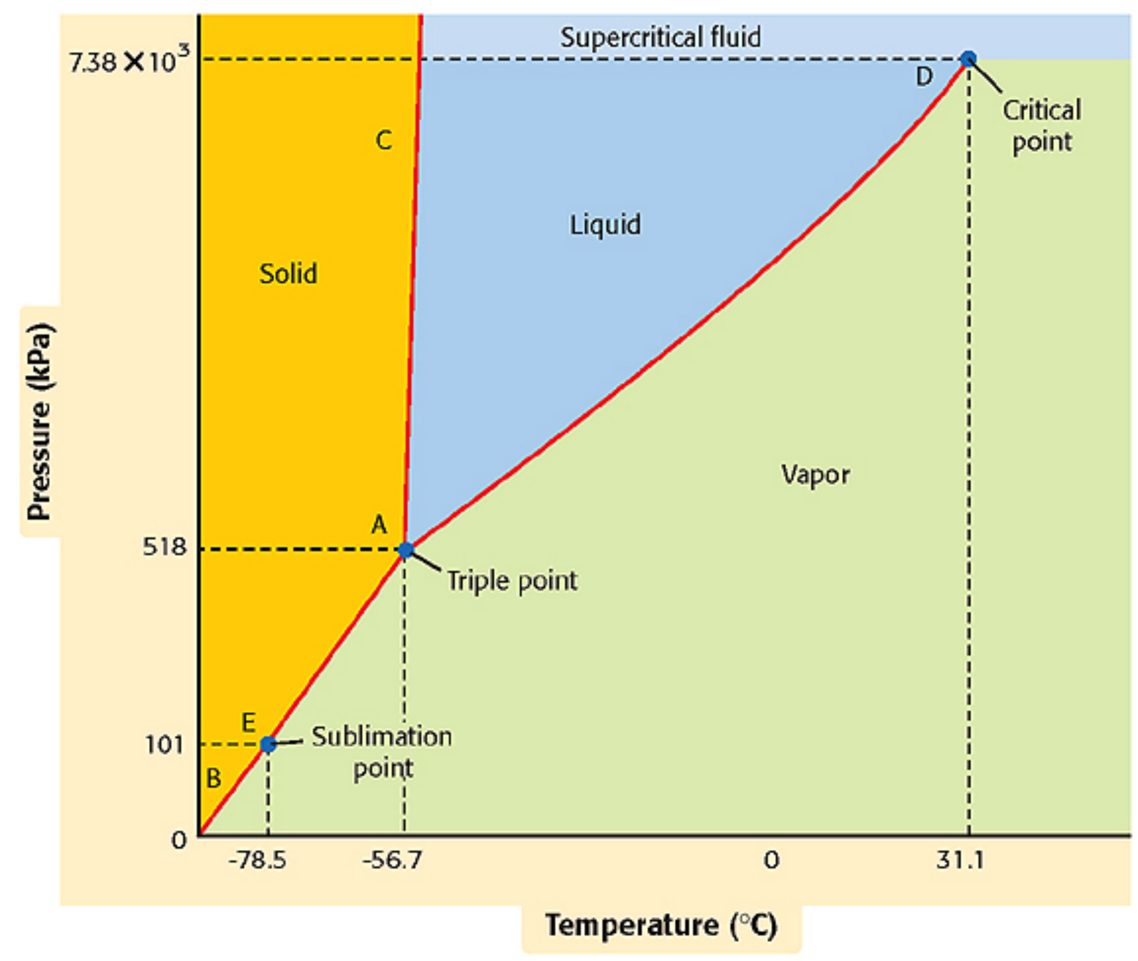
Click below to watch the Visual Concept.

Visual Concept

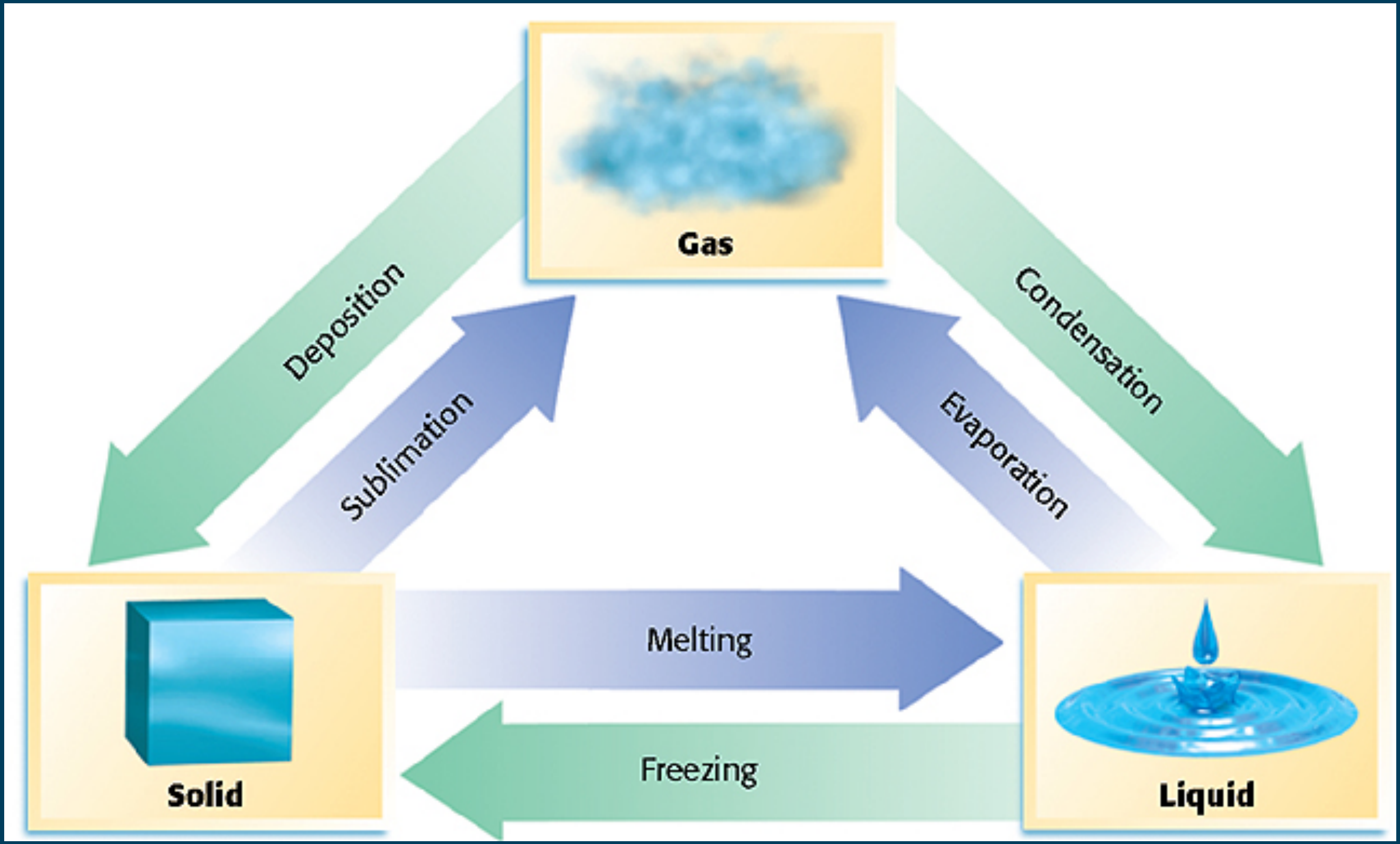
Phase Diagram for Water



Phase Diagram for CO2



Changes of State



Preview

- Lesson Starter
- Objectives
- Structure of Water
- Physical Properties of Water

Lesson Starter

- How would the water molecule's structure affect the properties of water?

- How will hydrogen bonding influence the properties of water?

Objectives

- **Describe** the structure of a water molecule.
- **Discuss** the physical properties of water. **Explain** how they are determined by the structure of water.
- **Calculate** the amount of energy absorbed or released when a quantity of water changes state.

Structure of Water

- Water molecules consist of two atoms of hydrogen and one atom of oxygen united by polar-covalent bonds.
- The molecules in solid or liquid water are linked by hydrogen bonding.
  - The number of linked molecules decreases with increasing temperature.
- Ice consists of water molecules in the hexagonal arrangement.

End Of Slide

Structure of a Water Molecule

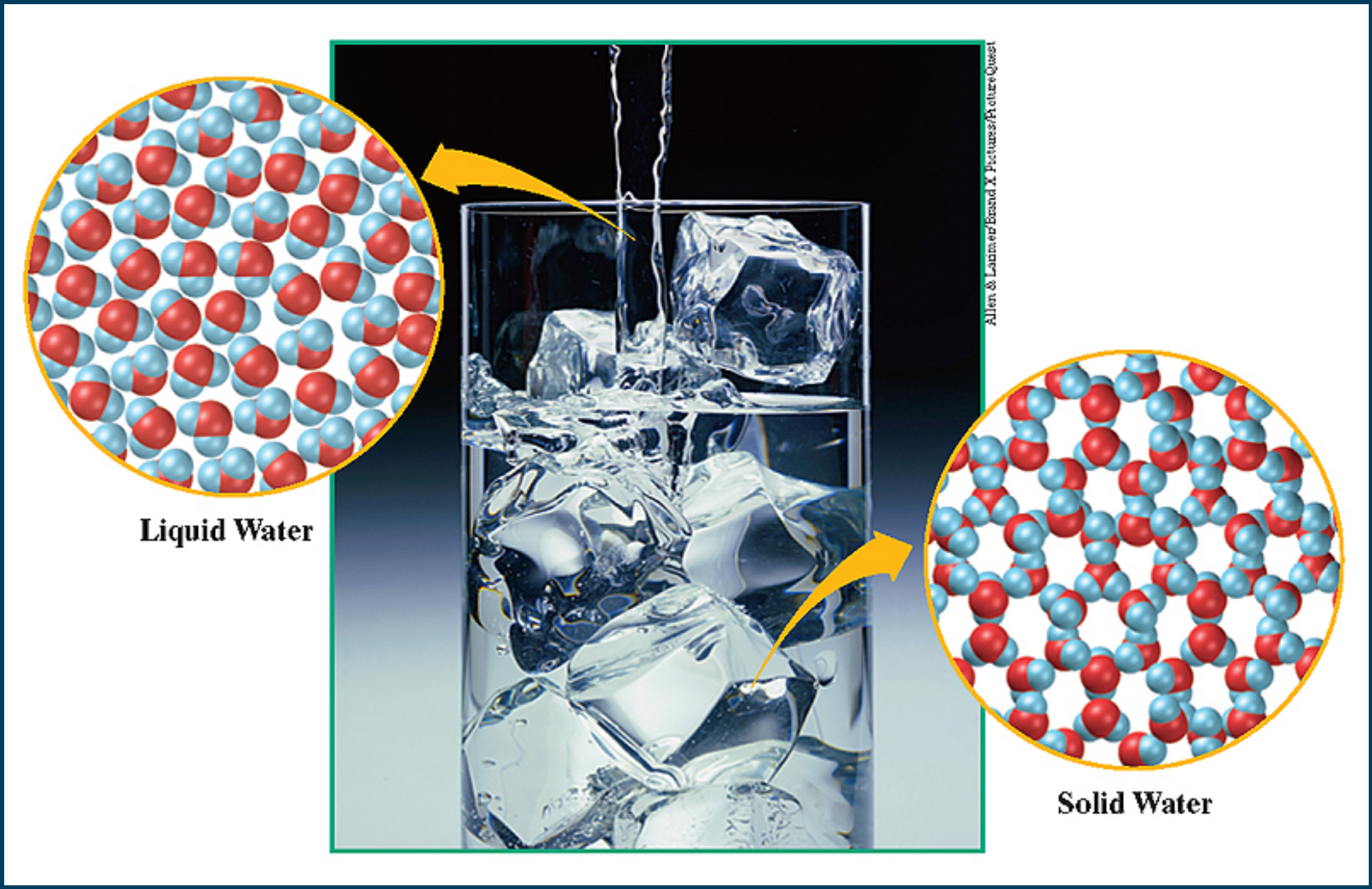
Click below to watch the Visual Concept.

Visual Concept

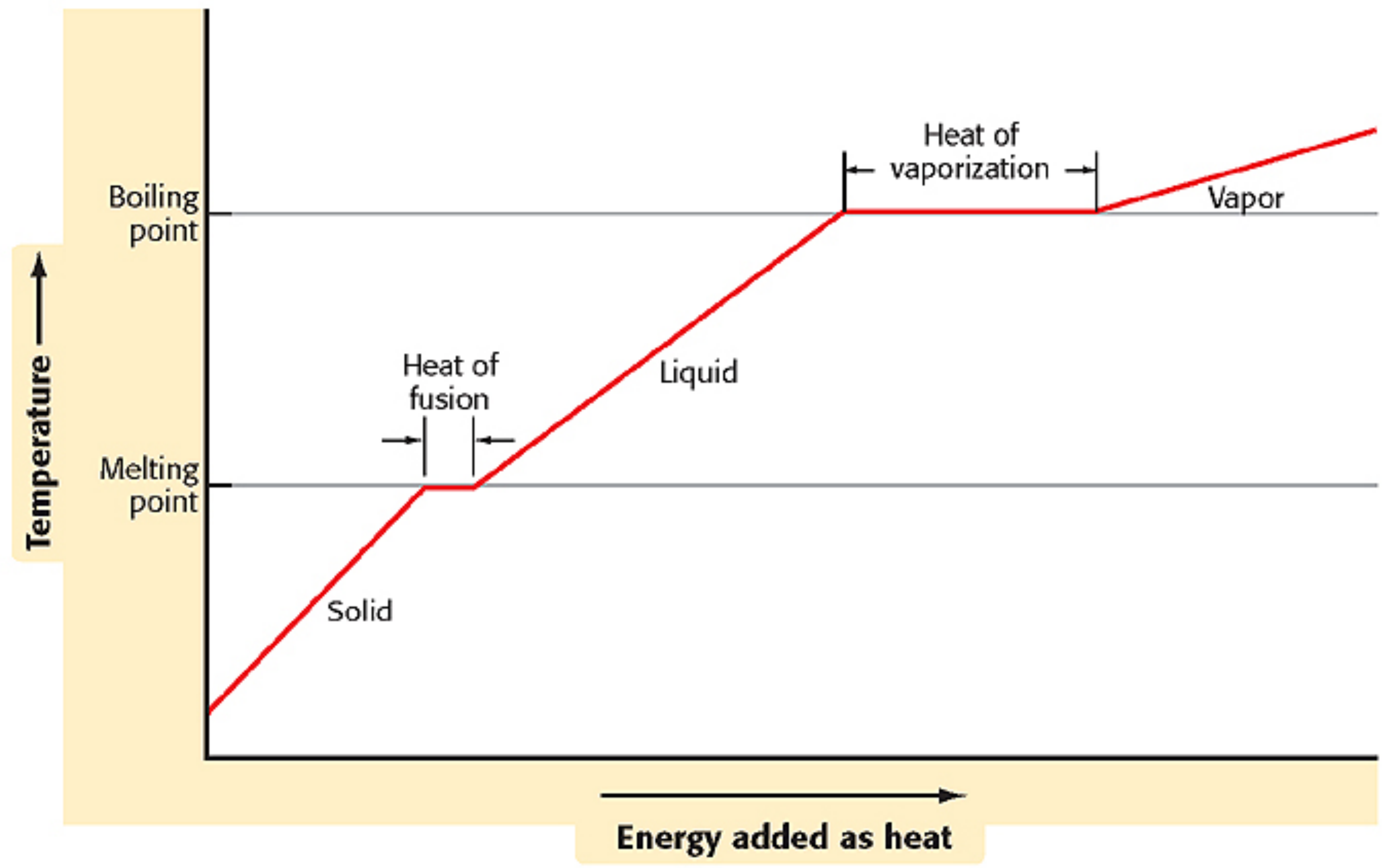
Structure of Water, *continued*

- The hydrogen bonds between molecules of liquid water at 0.°C are fewer and more disordered than those between molecules of ice at the same temperature.
  - Liquid water is denser than ice.
- As the temperature approaches the boiling point, groups of liquid water molecules absorb enough energy to break up into separate molecules.

Ice and Water



Heating Curve for Water



Physical Properties of Water

- At room temperature, pure liquid water is transparent, odorless, tasteless, and almost colorless.
- The molar enthalpy of fusion of ice is relatively large compared with the molar enthalpy of fusion of other solids.
- Water expands in volume as it freezes, because its molecules form an open rigid structure.
    - This lower density explains why ice floats in liquid water.

Physical Properties of Water, *continued*

- Both the boiling point and the molar enthalpy of vaporization of water are high compared with those of nonpolar substances of comparable molecular mass.
  - The values are high because of the strong hydrogen bonding that must be overcome for boiling to occur.
- Steam (vaporized water) stores a great deal of energy as heat.

End Of Slide

Physical Properties of Water, *continued*

Sample Problem A

How much energy is absorbed when 47.0 g of Ice melts at STP? How much energy is absorbed when this same mass of liquid water boils?

Physical Properties of Water, *continued*

Sample Problem A Solution

Given: mass of H2O(*s*) = 47.0 g;
mass of H2O(*l*) = 47.0 g;
molar enthalpy of fusion of ice = 6.009 kJ/mol;
molar enthalpy of vaporization = 40.79 kJ/mol

Unknown: energy absorbed when ice melts;
energy absorbed when liquid water boils

Solution:

- Convert the mass of water from grams to moles.

$$47.0 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} = 2.61 \text{ mol } H_2O$$

Physical Properties of Water, *continued*

Sample Problem A Solution, *continued*

- Use the molar enthalpy of fusion of a solid to calculate the amount of energy absorbed when the solid melts.
- Calculate the amount of energy absorbed when water boils by using the molar enthalpy of vaporization.

amount of substance (mol) × molar enthalpy of fusion or vaporization (kJ/mol)

= energy (kJ)

2.61 mol × 6.009 kJ/mol = 15.7 kJ (on melting)

2.61 mol × 40.79 kJ/mol = 106 kJ (on vaporizing or boiling)

End of Slide

End of Chapter 10 Show

< Back Next > Preview Main