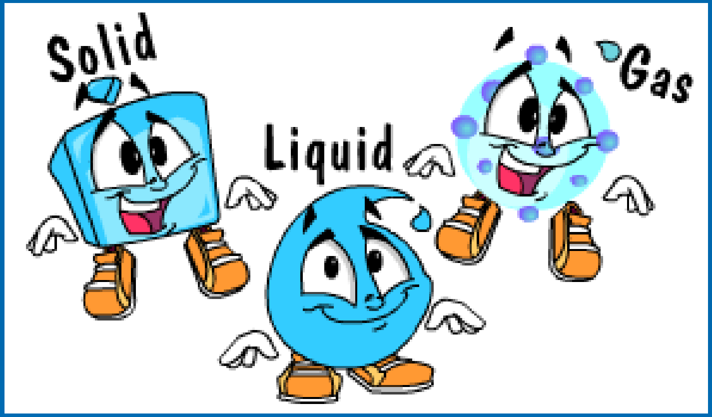
Unit 10: States of Matter (Chapter 13 and 15)

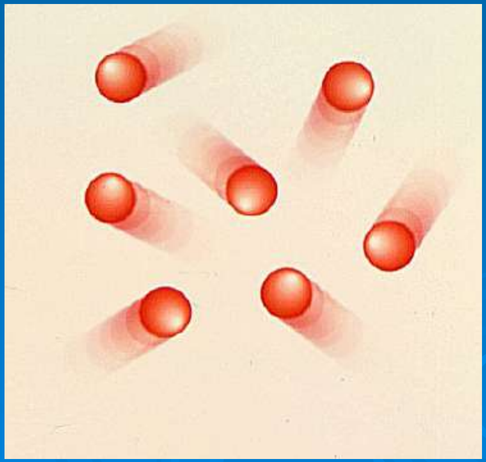


- Jennie L. Borders

Section 13.1 – The Nature of Gases

- Kinetic Energy is the energy of motion.

- The Kinetic Theory is based upon the idea that all matter is made of particles that are in constant motion.



The Kinetic Theory

- The particles of a gas are considered to be small, hard spheres with an insignificant volume.
- No attractive or repulsive forces exist between the particles.
- The motion of the particles in a gas is rapid, constant, and random.
- All collisions between particles in a gas are perfectly elastic.

The Kinetic Theory

- The particles of a gas travel in straight-line paths until they collide with another object.
- During an elastic collision, kinetic energy is transferred without loss from one particle to another, and the total kinetic energy remains constant.

Gas Pressure

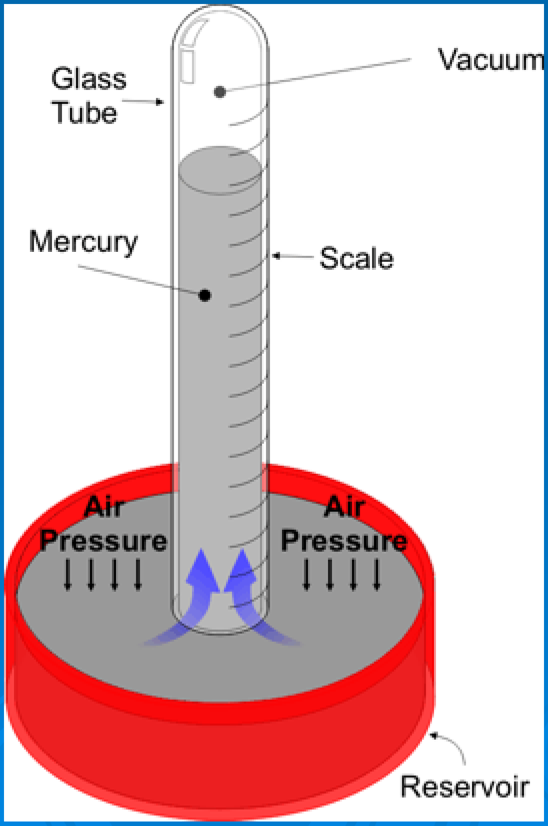
- Gas pressure is the force exerted by a gas per unit surface area of an object.
- Gas pressure is the result of simultaneous collisions of billions of rapidly moving particles in a gas with an object.
- An empty space with no particles and no pressure is called a vacuum.

Atmospheric Pressure

- Air exerts pressure on the earth because gravity holds the air particles in the Earth's atmosphere.
- Atmospheric pressure decreases as you climb a mountain because the density of Earth's atmosphere decreases as the elevation increases.

Barometer

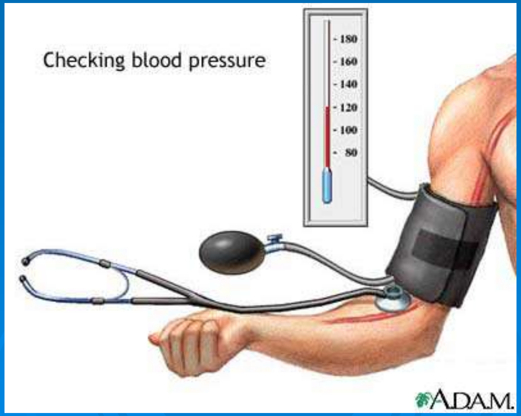
- A barometer is a device used to measure atmospheric pressure.



Units of Pressure

- The SI unit of pressure is the Pascal (Pa).
- The most common units of pressure are the atmosphere, millimeters of mercury, kilopascals, and torr.

1 atm = 760 mm Hg = 101.3 kPa = 760 Torr



Conversions of Pressure

Sample Problem 13.1

A pressure gauge records a pressure of 450 kPa. What is this measurement expressed in millimeters of mercury?

Answer:

$$450\ \text{kPa} \times \frac{760\ \text{mm Hg}}{101.3\ \text{kPa}} = \mathbf{3400\ mm\ Hg}$$

Conversion of Pressure

Practice Problem 1

What pressure in atmospheres does a gas exert at 385 mm Hg?

Answer:

$$385 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = \mathbf{0.51\ atm}$$

Kinetic Energy

- As a substance is heated, its particles absorb energy, some of which is stored within the particles.
- This increase in kinetic energy results in an increase in temperature.
- The particles in any substance at a given temperature have a wide range of kinetic energies.

Kinetic Energy

- Kinetic energy and Kelvin temperature are directly proportional.
- An increase in average kinetic energy causes the temperature to increase.
- A decrease in average kinetic energy causes the temperature to decrease.
- Absolute zero is the temperature at which the motion of particles theoretically stops.

Section Assessment

- Briefly describe the assumptions of the kinetic theory.
- How is the Kelvin temperature of a substance related to the average kinetic energy of its particles?
- Convert the following pressures to kilopascals.

a. 0.95 atmb. 45 mm Hg

Answers:
a. **96 kPa**b. **6.0 kPa**

Section 13.2 – The Nature of Liquids

- The high kinetic energy in gases and liquids allows the particles to flow past one another.
- Substances that can flow are called fluids.
- Intermolecular forces keep the particles in a liquid close together, which is why liquids have a definite volume,

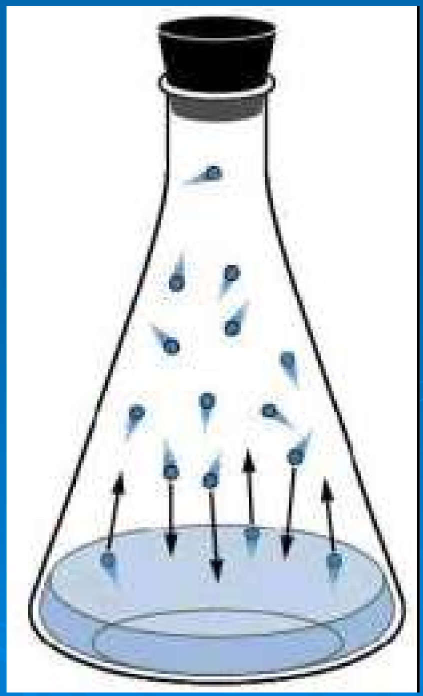
  unlike gases.

Evaporation

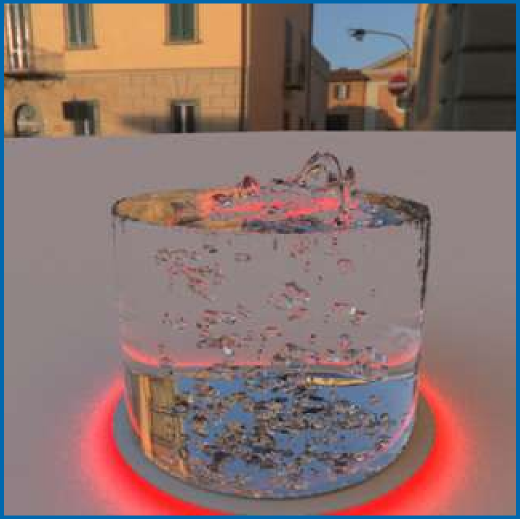
- The conversion of a liquid to a gas that is not boiling is referred to as evaporation.
- During evaporation, only molecules with the highest kinetic energy can escape the surface of a liquid.
- The particles left in the liquid have a lower average kinetic energy resulting in a lower temperature.

Vapor Pressure

- Vapor pressure is a measure of the force exerted by a gas above a liquid.
- An increase in temperature increases the vapor pressure produced by a liquid.



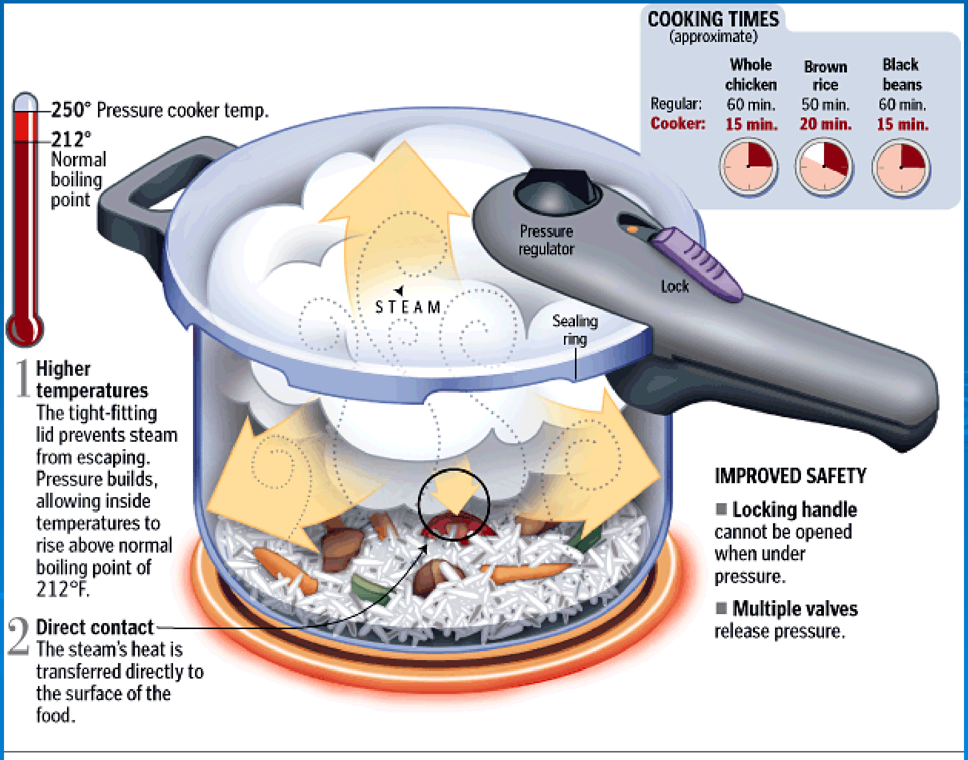
Boiling Point



- The rate of evaporation increases as the temperature increases.
- When a liquid is heated to a temperature at which particles throughout the liquid have enough kinetic energy to vaporize, the liquid begins to boil.
- The temperature at which the vapor pressure of the liquid is equal to the external pressure on the liquid is the boiling point.

Boiling and Pressure

- Because atmospheric pressure is lower at higher altitudes, boiling points decrease at higher altitudes.
- At higher external pressure, the boiling point increases.



Section Assessment

- In terms of kinetic energy, explain how a molecule in a liquid evaporates.
- Explain why the boiling point of a liquid varies with atmospheric pressure.
- Explain how evaporation lowers the temperature of a liquid.

Section 13.3 – The Nature of Solids

- The general properties of solids reflect the orderly arrangement of their particles and the fixed locations of their particles.
- When you heat a solid, its particles vibrate more rapidly as their kinetic energy increases.
- The melting point is the temperature at which a solid changes into a liquid.

Crystals

- In a crystal the particles are arranged in an orderly, repeating, three-dimensional pattern called a crystal lattice.
- The smallest group of particles within a crystal that retains the geometric shape of the crystal is known as the unit cell.

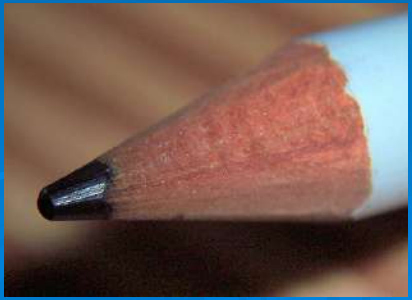



Melting

- Ionic solids have high melting points (above 300°C).
- Molecular solids have low melting points (below 300°C).
- Not all solids melt; some just decompose. (Ex. Wood)

Allotropes

- Allotropes are two or more different molecular forms of the same element in the same physical state.
- A common example is carbon: diamond, graphite, and bucky ball.
- Other examples include phosphorus, sulfur, and oxygen.



Non-Crystalline Solids

- An amorphous solid lacks an ordered internal structure.
- Examples include rubber, plastic, and asphalt.
- Glass is an amorphous solid that is a supercooled liquid. Glass is formed by cooling a liquid into a rigid state without crystallizing.

Section Assessment

- In general, how are the particles arranged in solids?
- How do allotropes of an element differ?
- How do the melting points of ionic solids generally compare with those of molecular solids?

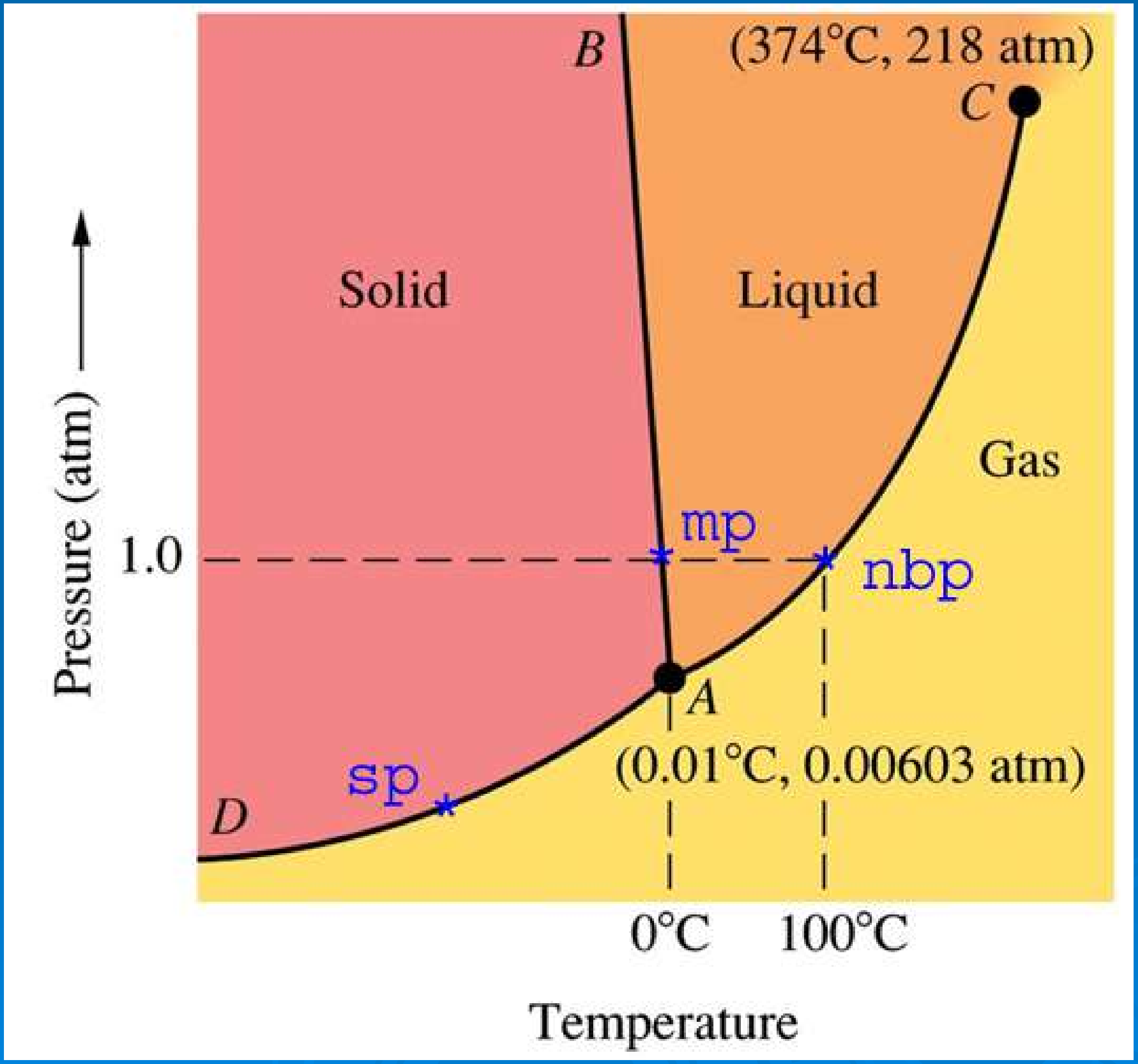
Section 13.4 – Changes of State

- The change of a substance from a solid to a vapor without passing through the liquid state is called sublimation.
- Dry ice and iodine are two examples of solids that sublimate.

Phase Diagrams

- The relationship among the solid, liquid, and gaseous states of a substance can be represented in a single graph called a phase diagram.
- The lines on a phase diagram indicate the conditions at which two phases occur in equilibrium.
- The triple point describes the only set of conditions at which all three phases occur in equilibrium.

Phase Diagram of Water

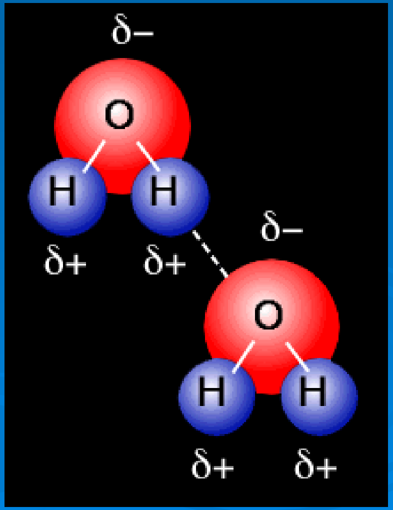


Section Assessment

- What does the triple point on a phase diagram describe?

Section 15.1 – Water and Its Properties

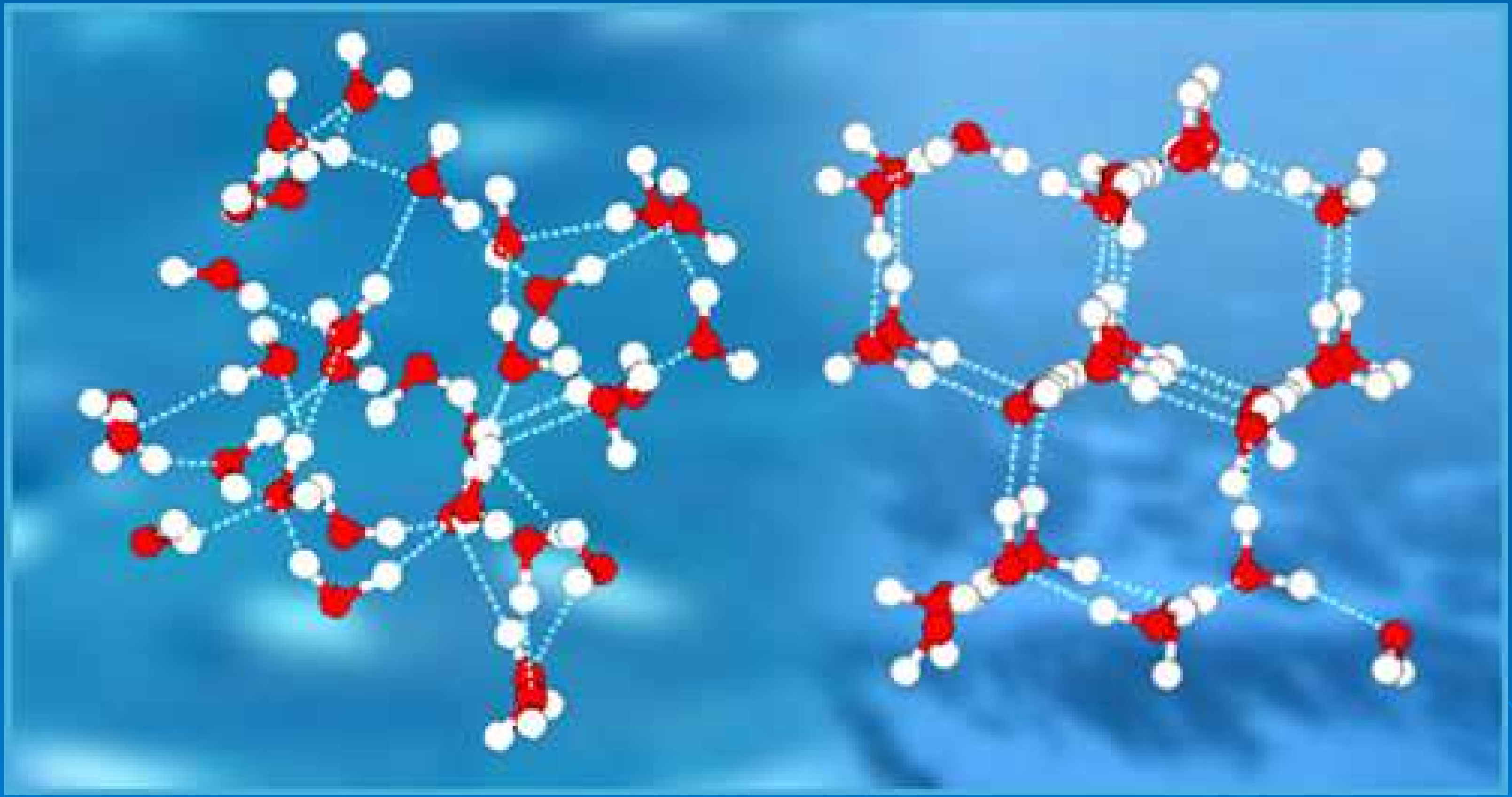
- A water molecule has a dipole moment because the oxygen is much more electronegative than the hydrogens.
- This strong dipole moment causes water molecules to have strong attractions for each other. These attractions are called hydrogen bonding.
- Hydrogen bonding describes many of the properties of water such as surface tension and vapor pressure.



Ice and Liquid Water

- Water is one of the few substances in which the solid state is less dense than the liquid state.
- This is the reason that ice floats in water.
- The structure of ice is a regular open framework of water molecules arranged like a honeycomb.
- When ice melts, the framework collapses and the water molecules pack close together, making the liquid more dense than the ice.

Ice and Liquid Water



Section Assessment

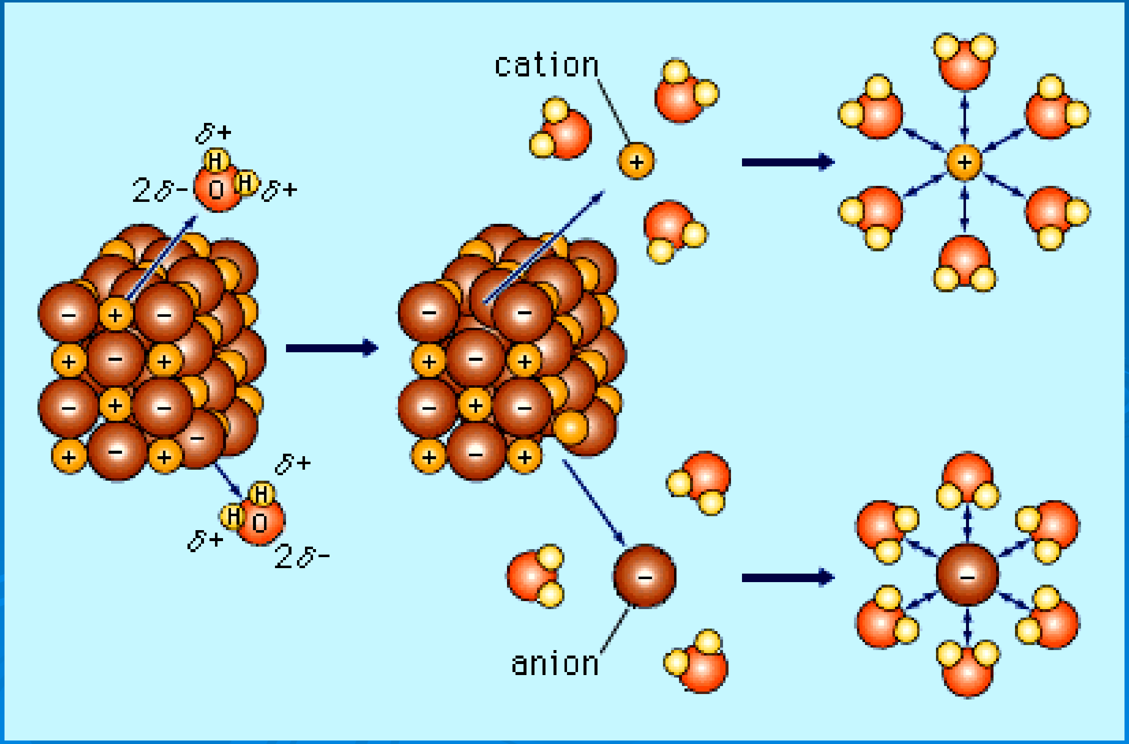
- What causes the high surface tension and low vapor pressure of water?
- How would you describe the structure of ice?

Section 15.2 – Homogeneous Aqueous Systems

- An aqueous solution is water that contains dissolved substances.
- In a solution, the dissolving medium is the solvent, and the dissolved particles are the solute.
- A solvent dissolves a solute.

Dissolving Ionic Solids

- As individual solute ions break away from a crystal, the negatively and positively charged ions become surrounded by solvent molecules and the ionic crystal dissolves.



Dissolution Rule

- As a rule, polar solvents such a water dissolve polar solutes such as ethanol.
- As a rule, nonpolar solvents such a gasoline dissolve nonpolar solutes such as oil.
- This relationship can be summed up in the expression "like dissolves like."

Electrolytes

- An electrolyte is a compound that conducts an electric current when it is in an aqueous solution or in the molten state.
- All ionic compounds are electrolytes because they dissolve into ions.
- A strong electrolyte fully breaks into ions.
- A weak electrolyte only partially breaks into ions.

Nonelectrolyte

- A substance that does not conduct electricity is a nonelectrolyte.
- Some polar compounds are nonelectrolytes in a pure state but become electrolytes when dissolved in water.

Hydrates

- A compound that contains water is called a hydrate.
- In writing the formula of a hydrate, use a dot to connect the formula of the compound and the number of water molecules per formula unit.
- Example:

  $$\mathbf{CuSO_4 \cdot 5H_2O}$$

Section Assessment

- In the formation of a solution, how does the solvent differ from the solute?
- Describe what happens to the solute and the solvent when an ionic compounds dissolves in water.
- Why are all ionic compounds electrolytes?
- How do you write the formula for a hydrate?
- Which of the following substances dissolve to a significant extent in water?

a. CH_4b. KClc. He

d. $MgSO_4$e. sucrosef. $NaHCO_3$

Section 15.3 – Heterogeneous Aqueous Systems

- A suspension is a mixture from which particles settle out upon standing because the solute particles are very large.
- An example is Italian salad dressing.

Colloids

- A colloid is a heterogeneous mixture containing particles that are smaller than a suspension but larger than a solution.
- A colloid's particles do not settle out with time.
- A colloid's particles are too small to be separated by filtering.
- Examples include whipped cream, milk, and Jell-O.

Section Assessment

- How does a suspension differ from a solution?
- What distinguishes a colloid from a suspension and a solution?
- Could you separate a colloid by filtering?

The End

FINALLY!!