

Chapter 10: Liquids and Solids

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*Liquids and solids show many similarities and are strikingly different from their gaseous state.

10.1 Intermolecular Forces

Intermolecular Forces-weak interactions between molecules composed of covalent and ionic bondings.

Dipole-dipole attractions-molecules with dipole moments line up with one another electrostatically where the positive and negative ends are close to each other

Strong Dipoles:

H-N	*increase in polarity going down from H-N to H-O
H-F	*have very strong dipole-dipole interactions called hydrogen bonding
H-O	

London dispersion forces-forces that exist among noble gases and non-polar molecules; an instantaneous pole that occurs accidentally in an atom and can induce a similar dipole in a neighboring atom (weak and short-lived).

*London dispersion forces increase as the size of the atom increases (increase chance of dipolar distribution). Example: Alkanes

10.2 The Liquid State

Surface tension-the resistance of a liquid to an increase in its surface area (polar molecules tend to have high surface tension)

Capillary action-the spontaneous rising of a liquid in a narrow tube due to cohesive (intermolecular forces) and adhesive (attractive forces) forces (adhesive > cohesive)

Viscosity-a measure of a liquid's resistance to flow; molecular complexity and polarity has to do with higher viscosity

10.3 An Introduction to Structures and Types of Solids

Categories of Solids:

Crystalline solids-highly regular arrangements of their components

Amorphous solids-considerable disorder in their structures

Lattice-3-dimensional system of points where each represents the position of the components (atoms, ions, molecules) that makes up the crystal

Unit cell-repeating unit of the lattice; x-ray diffraction is commonly used to determine the crystalline structure of solids

Three types of solids:

Ionic solid-a solid containing cations and anions that dissolves in water to give a solution containing the separated ions which are mobile and thus free to conduct electrical current.

Example: NaCl

Molecular solid-a solid composed of neutral molecules at the lattice points.

Example: C₁₂H₂₂O₁₂

Atomic solid-a solid that contains atoms at the lattice points

Example: C

10.4 Structure and Bonding in Metals

*Metals are characterized by having great electrical and thermal conductivity, malleability, and ductility.

*Metal atoms are packed tightly to minimize wasted space and increase interaction

Closest packing-an arrangement in crystals where the spheres are packed in layers, in which each sphere is surrounded by six others.

Hexagonal closest packed (hcp) structure-spheres are closest packed so that the spheres in the third layer are directly over those in the first layer.

Cubic closest packed (ccp) structure-the spheres in every fourth layer occupy the same vertical position each sphere has 12 equivalent nearest neighbors: 6 in the same layer, 3 in the layer above, 3 in the layer below

Bonding Models for Metals:

*Metal atoms are held together by strong forces, but they are relatively easy to move.

Electron sea model-a regular array of metal cations in a "sea" of valence electrons

Band model or Molecular orbital (MO) model-electrons travel around the crystal in molecular orbitals formed by the valence atomic orbitals of the metal atoms

*Metals conduct heat and electricity efficiently because the electrons are highly mobile

Metal Alloys:

*An alloy is a substance that is a mixture of elements with metallic properties

Substitutional alloys-an alloy where some of the host metal atoms are replaced by other metal atoms of similar size

Example: sterling silver, pewter, plumber's solder

Interstitial alloy-formed when some of the holes in the closest packed metal structure are occupied by small atoms

Example: steel

10.5 Carbon and Silicon: Network Atomic Solids

Network solids-atomic solids that contain strong directional covalent bonds and don't conduct heat and electricity efficiently (giant molecules)

Example: carbon and silicon

Carbon:

diamond, sp^3 hybridized

graphite, sp^2 hybridized

Silicon:

Silica-fundamental silicon-oxygen compound, which has the empirical formula SiO_2

Glass-an amorphous solid yielded by melted silicon (BrO_3 can be added to make borosilicate glass)

Ceramics-hardened clay, which contain silicates that are strong, brittle, and resistant to heat and chemicals

Semiconductors-conduct electricity better at higher temperatures as opposed to lower temperatures in metals (used in AC/DC converters)

10.6 Molecular Solids

Molecular solids-solids with discrete molecular units at each lattice position

*Forces that hold these solids together include London Dispersion Forces i.e. (CO_2 , I_2 , P_4 , and S_8) and dipole-dipole interactions i.e. (H_2O with hydrogen bondings)

10.7 Ionic Solids

Ionic solids-held together by strong electrostatic forces between two oppositely charged ions

*Packing is done in a way to maximize the electrostatic attractions and minimize repulsions of the oppositely charged ions (sample 10.4)

10.8 Vapor Pressure and Changes of State

Enthalpy of vaporization (ΔH_{vap})-the energy required to vaporize 1 mole of a liquid at a pressure of 1 atm

Dynamic equilibrium-rate of vaporization is equal to that of condensation

Vapor pressure-pressure of vapor at equilibrium

*Liquids that have high vapor pressure are said to be volatile

*Vapor pressure is principally determined by the substance's intermolecular forces

*Vapor pressure increases significantly with temperature

To calculate ΔH_{vap} :

$$\ln(P^{\text{vap}}_{T_1}/P^{\text{vap}}_{T_2}) = \Delta H_{\text{vap}}/R (1/T_2 - 1/T_1)$$

*Sample 10.6

Sublimation-a process in which a substance goes directly from the solid to the gaseous state (solids have a vapor pressure)

Changes of State:

ΔH_{fus} -enthalpy of fusion; the enthalpy change that occurs at the melting point when a solid melts (endothermic process)

Normal melting point-temperature at which the solid and liquid states have the same vapor pressure and the total pressure is 1 atm

Normal boiling point-the temperature at which the vapor pressure is 1 atm

Supercooling-a substance cooled below its freezing point without forming a crystalline solid

Superheating-this allows the vapor pressure of a liquid to be greater than the atmospheric pressure; vapor is released and may result in bumping

10.9 Phase Diagrams

Phase diagrams-shows the phases of a substance as a function of temperature and pressure

Triple point-all three states of matter exist at this temperature and pressure

Critical temperature-temperature above which the vapor cannot be liquefied no matter what pressure is applied

Critical pressure-pressure required to produce liquefaction at the critical temperature

Critical point-defined by T_c and P_c

*It would be helpful for students to read experiments 1 to 4 on pages 484-485

Applications of the Phase Diagram for Water:

*Solid/liquid boundary line has a negative slope, which means that the density of ice is less than liquid water

*Increased pressure decreases melting point

*Water reduces its volume with increased pressure by changing into a liquid (most condensed state)