Chapter 10: Liquids and Solids

Section 10.1 - Intermolecular Forces

Intramolecular bonding - when atoms form stable units called molecules by sharing electrons.

Intermolecular forces - weak interactions between molecules which can cause a **condensed state** of matter (liquids and solids). It is important to note that in this condensed state the molecule remains intact and retains its properties.

1. **Dipole-dipole forces** - molecules with dipole moments can attract each other electrostatically by lining up so that the positive and negative ends are close to each other. This is called a **dipole-dipole attraction**. These forces are typically only 1 % as strong as covalent or ionic bonds. **See Figure 10.2 p. 445**
2. **Hydrogen bonding** - particularly strong dipole-dipole forces, which are seen among molecules in which hydrogen is bound to a highly electronegative atom, such as nitrogen, oxygen, or fluorine. There are two reasons for being strong:
   a. The great polarity of the bond.
   b. The close approach of the dipoles, allowed by the very small size of the hydrogen atom.
   See Figure 10.3 p. 445
3. **London dispersion forces** - reaction between molecules which do not have dipole moments to allow them to be in a condensed state. As the electrons move about the nucleus, a momentary nonsymmetrical electron distribution can develop that produces a temporary dipolar arrangement of charge. This then affects the electron distribution of a neighboring atom.
   Polarizability - indicates the ease in which the electron "cloud" of an atom can be distorted to give a dipolar charge distribution. As the atomic number increases, the number of electrons increases, and there is an increased chance of the occurrence of momentary dipole interactions.
   See Figure 10.5 p. 447

Section 10.2 - The Liquid State

Surface tension - the resistance of a liquid to an increase in its surface area.

Capillary action - the spontaneous rising of a liquid in a narrow tube. This is due to two forces, **cohesive forces**, the intermolecular forces among the molecules of the liquid; and **adhesive forces**, the forces between the liquid molecules and their container.

Viscosity - a measure of a liquid's resistance to flow.

Section 10.3 - An Introduction to Structures and Type of Solids

Crystalline solids - solids with a highly regular arrangement of their components.

Amorphous solids - solids with considerable disorder in their structures.

Lattice - a three dimensional system of points designating the positions of the components (atoms, ions, or molecules) that make up the substance.

Unit cell - the smallest repeating unit of a lattice. see fig. 10.9

X-ray diffraction - most common way of determining the structures of crystalline solids.

Types of crystalline solids - **see fig. 10.12 p.454**

1. **ionic solids** - have ions at the points of the lattice that describes the structure of the solid
2. **molecular solids** - has discrete covalently bonded molecules at each of its lattice points.
3. **atomic solids** - have atoms at the lattice points that describe the structure of the solid.
   a. **metallic** - specialized type of delocalized non directional covalent bonding occurs.
   b. **network** - atoms bond together with strong directional covalent bonds which leads to giant molecules.
   c. **group 8A** - noble gas elements are attracted to each other with London dispersion forces.

Section 10.4 - Structure and Bonding in Metals

Properties of metals, such as high thermal and electrical conductivity, malleability and ductility, can be traced to non directional covalent bonding found in metallic crystals.

Closest packing - a model which pictures a metallic crystal as containing spherical atoms packed together and bonded to each other in all directions, in a manner that most efficiently uses the available space. **See Figure 10.13 p. 455**

1. **Hexagonal closest packed (hcp) structure** - an aba arrangement (the atoms in the first layer lie directly above the atoms in the third layer) has a hexagonal unit cell. **See fig 10.14 p. 456**
2. **Cubic closest packed (ccp) structure** - an abc arrangement (the atoms in the first layer do not lie directly above the atoms in the third layer) has a face-centered cubic unit cell. **See fig 10.15 p. 455**

Bonding in most metals is both strong and non directional. That is, although it is difficult to separate metal atoms, it is relatively easy to move them, provided the atoms stay in contact with each other.

Molecular orbital (MO) (or band model) - in this model, the electrons are assumed to travel around the metal crystal in molecular orbitals formed from the valence atomic orbitals of the metal atoms. When many metal atoms interact, the large number of resulting molecular orbitals become more closely spaced and finally form a virtual continuum of levels, called bands. **See fig 10.19 and 10.20**

Alloy - a substance that contains a mixture of elements and has metallic properties.

1. **Substitutional alloy** - some of the host metal atoms are replaced by other metal atoms of similar size.
2. **Interstitial alloy** - formed when some of the interstices (holes) in the closest packed metal structure are occupied by small atoms.

Section 10.5 - 10.7 - optional reading

Section 10.8 - Vapor Pressure and Change of State

Evaporation or Vaporization - when molecules of a liquid escape the liquid's surface and form a gas. It is an endothermic process because energy is required to overcome relatively strong intermolecular forces in the liquid.

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

Condensation - the process by which vapor molecules reform a liquid.

Equilibrium - the point at which no further net change occurs in the amount of liquid or vapor because the two opposite processes exactly balance each other. A system at equilibrium is dynamic (in constant motion) on the molecular level but shows no macroscopic changes. **See fig 10.36 and 10.37 p. 475**

Vapor pressure - the pressure of the vapor present at equilibrium.

$$P_{vapor} = P_{atmosphere} - P_{Hg\ column}$$

The vapor pressure of a liquid is principally determined by the size of the intermolecular forces in the liquid. Liquids which have high vapor pressures are said to be volatile - they evaporate rapidly form an open dish and have low intermolecular forces. Liquids which have low vapor pressures have high intermolecular forces, and molecules need high energies to escape the vapor phase.

In general, substances with large molar masses have relatively low vapor pressures, mainly because of the large dispersion forces. The more electrons a substance has, the more polarizable it is, and the dispersion forces are greater.

An increase in temperature significantly increases vapor pressure. If the vapor pressure for a liquid is known for a given temperature, then a second vapor pressure can be found for a second temperature by using the Clausius-Clapeyron equation.

Solids also have vapor pressures - this leads to a process called **sublimation** - a process in which a substance goes directly form the solid to the gaseous state.

Heating curve - a plot of temperature versus time for a process where energy is added at a constant rate.

1. **Melting point** - plateau at which energy added is used to change the material from a solid to a liquid (there is no rise in overall temperature during the process). this is also the point at which the vapor pressure of the liquid and the solid are equal.
2. **Heat of fusion (or enthalpy of fusion) ΔH_{fus}** - the enthalpy change that occurs at the melting point when a solid melts.
3. **Boiling point** - plateau at which energy added is used to change the material from a liquid to a gas (there is no rise in overall temperature during this process).
4. **Normal melting point** - the temperature at which the solid and liquid stats have the same vapor pressure under conditions where the total pressure is 1 atmosphere.
5. **Normal boiling point** - the temperature at which the vapor pressure of the liquid is exactly 1 atmosphere. Boiling occurs when the vapor pressure of a liquid becomes equal to the pressure of its environment.

Section 10.9 - Phase Diagrams

Phase diagram - a convenient way of representing the phases of a substance as a function of temperature and pressure.

Triple point - solid and liquid have identical vapor pressures, and all three states of matter coexist in a closed system.

Critical temperature - the temperature above which the vapor can not be liquefied no matter what pressure is applied.

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

Critical point - the combination of the critical pressure and the critical temperature.

****Notes have been derived from Zumdahl 4th ed. - All page and table references are made to this edition.**