Chapter 6: States of Matter: Gases, Liquids, and Solids

6.1 The Gaseous State

- **Ideal Gas Concept**
  - **Ideal gas** - a model of the way that particles of a gas behave at the microscopic level.
  - We can measure the following of a gas:
    - temperature,
    - volume,
    - pressure and
    - quantity (mass)
  - We can systematically change one of the properties and see the effect on each of the others.

- **Measurement of Gases**
  - The most important **Gas Laws** involve the relationship between
    - number of moles (n) of gas
    - volume (V)
    - temperature (T)
    - pressure (P)
  - **Pressure** - force per unit area.
  - Gas pressure is a result of force exerted by the collision of particles with the walls of the container.
  - **Barometer** - measures atmospheric pressure.
  - A commonly used unit of pressure is the atmosphere (atm).
  - 1 atm is equal to: 760 mmHg = 760 torr = 76 cmHg

- **Boyle's Law** - <u>volume</u> of a gas is *inversely proportional* to <u>pressure</u> if the temperature and number of moles is held constant.

- **Charles' Law** - <u>volume</u> of a gas varies *directly* with the <u>absolute temperature</u> (K) if pressure and number of moles of gas are constant.

- **Temperature-Pressure Relationship at Constant Volume -** The pressure of a gas is directly proportional to the Kelvin temperature if the volume is kept constant.

- **Combined Gas Law -** This law is used when a sample of gas undergoes change involving volume, pressure, and temperature simultaneously.

- **Avogadro's Law** - equal volumes of an ideal gas contain the same number of moles if measured under the same conditions of temperature and pressure.

- **Molar Volume** - the volume occupied by 1 mol of any gas determined by Avogadro. At STP the molar volume of a gas is 22.4 L.
  - **STP** - Standard Temperature and Pressure
    - T = 273 K (or 0°C)
    - P = 1 atm
    - We will learn to calculate the volume later.

- **The Ideal Gas Law -** Combining Boyle's Law, Charles' Law, T-P Relationship Law and Avogadro's Law gives the Ideal Gas Law.

- **Dalton's Law of Partial Pressures -** a mixture of gases exerts a pressure that is the sum of the pressures that each gas would exert if it were present alone under the same conditions.
  - For example, the total pressure of our atmosphere is equal to the sum of the pressures of N_2 and O_2.

- **Kinetic Molecular Theory of Gases**
  - 1. Gases are made up of small atoms or molecules that are in constant and random motion.
  - 2. The distance of separation is very large compared to the size of the atoms or molecules.
    - The gas is mostly empty space.
  - 3. All gas particles behave independently.
    - No attractive or repulsive forces exist between them.
  - 4. Gas particles collide with each other and with the walls of the container without losing energy.
    - The energy is transferred from one atom or molecule to another.
  - 5. The average kinetic energy of the atoms or molecules is proportional to absolute temperature.
    - K.E. = $1/2mv^2$ so as temperature goes up, the speed of the particles goes up.
  - How does the Kinetic Molecular Theory of Gases explain the following statements?
    - Gases are easily compressible.
    - Gases will expand to fill any available volume.
    - Gases have low density.
  - Remember: pressure is a force per unit area resulting from collision of gas particles with the walls of the container. If pressure remains constant why does volume increase with temperature?
  - Gases behave most ideally at low pressure and high temperatures.

- **Ideal Gases Vs. Real Gases**
  - In reality there is no such thing as an ideal gas. Instead this is a useful model to explain gas behavior.
  - Non-polar gases behave more ideally than polar gases because attractive forces are present in polar gases.

6.2 The Liquid State

- Liquids are practically incompressible.
  - Enables brake fluid to work in your car
- **Viscosity** - a measure of a liquids resistance to flow.
  - Flow occurs because the molecules can easily slide past each other.
  - Glycerol - example of a very viscous liquid.
  - Viscosity decreases with increased temperature.
- **Surface Tension** – a measure of the attractive forces exerted among molecules at the surface of a liquid.
- Surface molecules are surrounded and attracted by fewer liquid molecules than those below.
- Net attractive forces on surface molecules pull them downward.
  - Results in "beading"
- **Surfactant** - substance added which decreases the surface tension
  - example: soap
- Kinetic Theory - Liquid molecules are in continuous motion, with their *average* kinetic energy directly proportional to the Kelvin temperature.
- What happens when you put water in a sealed container?
- Both liquid water and water vapor will exist in the container. How does this happen below the boiling point?
- Once there are molecules in the vapor phase, they can be converted back to the liquid phase
- **Evaporation** - the process of conversion of liquid to gas, at a temperature too low to boil
- **Condensation** - conversion of the gas to the liquid state.
- When the rate of evaporation equals the rate of condensation, the system is at equilibrium.
- Vapor pressure of a liquid - the pressure exerted by the vapor at equilibrium
- **Boiling point** - the temperature at which the vapor pressure of the liquid becomes equal to the atmospheric pressure.
- **Normal boiling point** - temperature at which the vapor pressure of the liquid is equal to 1 atm.
- What happens when you go to a mountain where the atmospheric pressure is lower than 1 atm? The boiling point lowers.
- Boiling point is dependant on the inter-molecular forces
  - Polar molecules have higher b.p. than nonpolar molecules.
- **Van der Waals Forces** - are types of intermolecular forces.
- Consists of:
  - Dipole-dipole interactions (section 4.5)
  - London forces
- London Forces:
  - Exist between all molecules
  - Is the only attractive force between nonpolar atoms or molecules

- Electrons are in constant motion. This can create temporary dipoles among atoms. These dipoles can interact to cause attraction.
- Hydrogen bonding:
  - not considered a Van der Waals Force
  - is a special type of dipole-dipole attraction
  - is a very strong intermolecular attraction causing higher than expected b.p. and m.p.
- Requirement for hydrogen bonding:
  - molecules have H directly bonded to O, N or F

6.3 The Solid State

- Particles highly organized and well defined fashion
- Fixed shape and volume
- Properties of Solids:
  - incompressible
  - m.p. depends on strength of attractive force between particles
  - **Crystalline solid** - regular repeating structure
  - **Amorphous solid** - no organized structure.
- **Types of Crystalline Solids**
  - 1. Ionic Solids
    - held together by electrostatic forces
    - high m.p. and b.p.
    - hard and brittle
    - if dissolves in water, electrolytes
    - NaCl
  - 2. Covalent Solid
    - Held together entirely by covalent bonds
    - high m.p. and b.p.
    - extremely hard
    - Diamond
  - 3. Molecular solids
    - molecules are held together with intermolecular forces
    - often soft
    - low m.p.
    - often volatile
    - ice
  - 4. Metallic solids
    - metal atoms held together with metal bonds
    - metal bonds
    - overlap of orbitals of metal atoms
    - overlap causes regions of high electron density where electrons are extremely mobile - conducts electricity