CHEMISTRY NOTES – Chapter 10

States of Matter

Goals: To gain an understanding of:

1. The kinetic theory of matter.
2. Avogadro's hypothesis.
3. The nature of liquids, solids and gases.
4. Phase changes.

NOTES:

The kinetic theory of matter states that particles which make up all types of matter are in constant motion.

The kinetic theory of gases assumes three things:

- A gas is composed of tiny particles, usually atoms or molecules that have negligible volume and virtually no attractive or repulsive forces exist between particles. Virtually all of the volume of a gas is empty space through which these particles move.
- These particles move in a random and rapid constant motion until they collide with other particles or obstacles.
- The collisions between particles are elastic, meaning the kinetic energy is conserved - the amount of motion before and after the collision is equal. To understand the idea of elasticity think of dropping a "super ball." The amount of motion is pretty much conserved as the ball bounces and can therefore be described as elastic. Now think about dropping a ball of clay - inelastic.

The three factors which determine the physical behavior of gases are:

- Pressure - pressure will compress a gas, reducing its volume and giving it a greater density and concentration of particles.
- Temperature - The higher the temperature, the greater the kinetic energy of the particles and vice versa.
- Volume - A change in the volume of a contained gas will change the gas' concentration and pressure.

Gas pressure is the result of the simultaneous collisions of a very large number of gas particles striking an object.

A vacuum is the relatively empty space that results from the removal of gas particles from a closed container.

Atmospheric pressure is the result of the simultaneous collisions of atmospheric molecules striking objects. Three units used to measure atmospheric pressure are

- pascals (Pa) - SI unit of pressure
- millimeter of mercury (mm Hg) - the pressure needed to support one mm Hg
- atmospheres (atm) one standard atmosphere is equal to 760 mm Hg at 25 ºC

Two types of barometers are:

- mercury barometer - consists of an evacuated glass cylinder immersed in a pool of mercury. The pressure of the atmosphere then pushes the mercury into the glass cylinder and then can be measured (usually in mm).
- aneroid barometer - most common today - has a diaphragm which will be pushed in with greater pressure and be allowed to push out with less air pressure. These changes are then transferred to a dial or pointer which indicates air pressure.
Kinetic energy is the energy of motion. $$KE = \frac{1}{2}mv^2$$

When heat energy is added to a gas its particles will increase in kinetic energy (move faster) and the internal energy of the particles will increase. Temperature is a measure of the average kinetic energy of the particles of a substance.

The two condensed states of matter are solids and liquids. Condensed means they do not significantly change in volume with changes in pressure. Gases are not condensed, their volume changes dramatically with changes in pressure.

**The Liquid State:**

The particles of liquids have a much greater attraction for each other than do gases. This gives liquids a much greater density and a definite volume. The attractive force is not so great to hold the particles rigidly in place so that liquids have the property of fluidity - they flow.

**Structural Model for Liquids**
1. Strong intermolecular forces (like solids)
2. Considerable molecular motion (like gases)

**Properties of Liquids**
1. Fluids: substances that can flow and therefore take the shape of their container
2. 10% less dense than solids (average) (water is an exception – its liquid state is less dense that its solid state)
3. more dense than gases
4. relatively incompressible

Vaporization, or evaporation, is a cooling process (endothermic). A substance in its gaseous state has more kinetic energy than in its liquid state so it must absorb energy (take energy away) from its environment (reducing its kinetic energy or temperature) to change from a liquid to a gas.

Condensation is the opposite, it is a warming process. When a gas condenses into its liquid state the particles lose energy which is absorbed by the environment causing it to gain energy. This is, in part, the cause of storms such as hurricanes. When warm, moist air rises over the southern Atlantic ocean the water vapor it carries condenses and releases heat energy. This heat energy draws up more warm, moist air and its water vapor condenses creating a positive feedback system which creates the hurricane. Hurricanes generally die soon after moving over land because the moisture supply which generated the heat by condensation is cut off.

Vapor pressure is the simultaneous collision of very many vapor particles with the walls of a sealed container. If the container were open the particles would simple escape into the atmosphere.

Dynamic equilibrium is a state in which the rate of vaporization equals the rate of condensation in a sealed container so that the volumes of liquid and vapor remain constant. So it is "dynamic" - constant change (liquid ⇋ gas) and at equilibrium - constant volumes of liquid and vapor.

The boiling point of a liquid is the temperature at which it boils at a certain pressure. At the boiling point the vapor pressure exerted upward by the vapor particles equals the downward pressure of the atmosphere ("pressure up = pressure down"). The normal boiling point of a liquid is the temperate at which it boils under one atmosphere of pressure.
Since the boiling point of a liquid is directly related to external pressure, the boiling point of a liquid will change with external pressure. If the air pressure is low, such as at a high altitude, it would require less vapor pressure and the boiling temperature would be less than the normal boiling point. For this reason food needs to be boiled for longer at higher altitude to become fully cooked.

Conversely, if food is put in a pressure cooker, the high pressure allows water to get much hotter than 100 °C and food will cook much quicker.

When water boils it is being converted from a liquid to a gas. This is an endothermic process as the gaseous state has more energy than the liquid state as explained in objective 19 above. The reason water boils at a certain temperature (even though heat from a stove maybe being still applied) is that the vaporizing particles take away enough energy to keep the average kinetic energy the same and thus the temperature remains the same.

The melting point of a substance is the temperature at which the substance changes from its solid state to its liquid state. The value of a substances melting point and freezing points are the same (the change of state is the same, only in different directions).

Solids have a still greater attractive force between particles so that solids, in general, have a greater density than liquids and have both a definite volume and a definite shape. The particles of solids still move according to the kinetic theory, but they for the most part will vibrate and rotate in place.

The Solid State

Properties of Solids
A. Types of Solids
   1. Crystalline Solids - substances in which the particles are arranged in an orderly, geometric, repeating pattern
      a. Ionic crystals
      b. Covalent network crystals
      c. Diamond, quartz
      d. Metallic crystals
      e. Covalent Molecular crystals
      f. Ice
   2. Amorphous Solids - substances in which the particles are arranged randomly (glass)
B. Definite Shape and Volume
C. Definite Melting Point
D. High Density and Incompressibility

Ionic compounds generally have high melting points and molecular compounds have low melting points.

Crystals are solids which have their particles (atoms, ions or molecules) arranged in a very definite, orderly, repeating three dimensional pattern.

A crystal lattice is a repeating array within a crystal.

A unit cell is the smallest group of particles within a crystal lattice that keeps the geometric shape of the crystal.

Crystalline substances are composed of crystals as described above. An example of a crystalline substance would be quartz, or diamond or ruby etc..

Amorphous ("a" = without, "morph" = shape) substances do not have a regular, internal, ordered structure. Examples of amorphous substances are plastic, rubber and glass.

Supercooled liquids are amorphous substances such as glass. They are inorganic materials that have solidified without crystallizing. I have used both the words "liquid" and "solid" in this description which makes this idea pretty vague and unclear, but if you do any searching on the web on this topic you will find that it is pretty unclear as to whether glass is a solid or a liquid.
A phase change is a change from one state of matter (solid, liquid or gas) to another. For pure substances the temperature remains constant as phase change occurs. The temperature will change while a substance is in a particular phase. For a phase change to occur or for a substance in a particular phase to change to have a change of temperature energy must be added or taken away. This can be observed in the following graph showing the phase changes of pure water at one atm of pressure. During melting and freezing the temperature is constant at 0 ºC (at 100 ºC for vaporization and condensation) but energy is being added or taken away. During the liquid phase both temperature and energy change from point to point along the line.

Sublimation is the process by which a solid substance changes directly into a gaseous substance without going through the liquid phase. the most common example of sublimation is dry ice. Dry ice is solid (frozen) carbon dioxide which sublimes into gaseous carbon dioxide.

Plasma is a state of matter in which gaseous atoms become excited to the point that electrons fly away from the atoms creating an ionized gas-like state.

Partially ionized plasmas occur at low temperatures and have fewer atoms ionized than highly ionized plasmas. examples of partially ionized plasmas include fluorescent lights, neon lights, and flames.

Highly ionized plasmas occur at much higher temperatures and have many to all of the atoms in the ionized state. An example of this type of plasma is the matter which makes up stars (primarily hydrogen nuclei fusing into helium nuclei).

Changes of State

Phase Changes
- Evaporation (liquid → gas) (below the boiling point)
- Vaporization (liquid → gas) (at the boiling point)
- Boiling (liquid → gas)
- Condensation (gas → liquid)
- Melting (fusion) (solid → liquid)
- Freezing (or Solidification) (liquid → solid)
- Sublimation (solid → gas)
- Deposition (gas → solid)

Equilibrium Vapor Pressure of a Liquid

A. Equilibrium Vapor Pressure
   1. The pressure exerted by a vapor in equilibrium with its corresponding liquid at a given temperature

B. Volatile Liquids
   1. Liquids that have weak forces of attraction and evaporate easily

C. Nonvolatile Liquids
   1. Liquids that have strong forces of attraction and do not evaporate easily
Boiling
A. Boiling
1. The conversion of a liquid to a vapor within the liquid as well as at its surface. It occurs when the equilibrium vapor pressure of the liquid equals the atmospheric pressure (or the external pressure)

B. Boiling Point
1. The temperature at which the equilibrium vapor pressure of the liquid equals the atmospheric pressure
   a. Water boils at 100 °C at 1 atm pressure (the NORMAL boiling point)
   b. Water boils above 100 °C at higher pressures
   c. Water boils below 100 °C at lower pressures
   d. Strong attractive forces between water particles (hydrogen bonding) results in a high heat of vaporization

Phase Diagrams
A. Phase Diagram
1. A graph of pressure versus temperature that shows the conditions under which the phases of a substance exist

B. Triple Point
1. The temperature and pressure conditions at which the solid, liquid, and vapor of the substance can coexist at equilibrium

C. Critical Temperature
1. The temperature at above which the substance cannot exist in the liquid state, regardless of pressure
   a. For water, the critical temperature is 373.99 °C

D. Critical Pressure
1. The lowest pressure at which the substance can exist as a liquid at the critical temperature
   a. For water, the critical pressure is 217.75 atm

E. Critical Point
1. The point on the graph describing simultaneously the critical temperature and the critical pressure
   a. For water P = 217.75 atm Temperature = 373.99 °C