14a The Nature of Liquids and Water

A Model for Liquids

- According to the kinetic theory, like gases, the particles in liquids are also moving (though not as quickly).
  - Because of this motion, liquids can flow like gases.
- However, unlike gases, the particles in liquids are attracted to each other by intermolecular forces.
  - These intermolecular forces hold the particles of a liquid together and therefore liquids have a definite volume.
  - Also, because of these intermolecular forces, the particles of a liquid are much closer together than the particles in a gas.
    - Liquids have almost no compressibility. (Increased pressure has almost no effect on the volume of a liquid.)

Evaporation

- **Vaporization** – the conversion of a liquid to a gas/vapor
  - Boiling
  - Evaporation – vaporization at the surface of a liquid that is not boiling
    - Some particles in the liquid have enough kinetic energy to overcome the intermolecular forces and break away from the surface of the liquid to become a gas.
    - When the liquid is heated, the kinetic energy increases and more particles are able to escape.
    - After these particles with a high kinetic energy escape, the remaining liquid has a lower average kinetic energy and is therefore cooler.
      - This is how sweating cools down your body!

Vapor Pressure

- If a liquid is in a closed container and begins to evaporate, these new gas particles collide with the container and cause pressure, called **vapor pressure**.
At the same time that some particles are evaporating into gas, some of the gas particles also slow down and turn back into a liquid. Eventually, these two things will balance out and the vapor pressure will be constant.

As you know, an increase in temperature causes more evaporation. Therefore, at higher temperatures, vapor pressure increases as more particles are turning into gas.

Boiling Point

When the temperature is heated high enough, a large number of the particles have enough kinetic energy to escape, not only from the surface. Bubbles of vapor form inside the liquid and the rise to the surface and escape into the air.

- This temperature is called the **boiling point**, and it is the temperature that the vapor pressure of the liquid is equal to the external pressure.

Boiling Point and Pressure Changes

- Remember that in order to boil, the vapor pressure must be equal to the external pressure on the liquid. Therefore, if the external pressure decreases (like up in the mountains), the boiling point will also be lower, and if the external pressure increases (like in a pressure cooker), the boiling point will be higher.

- Because a liquid can have multiple boiling points depending on the external pressure, we often talk about something called the **normal boiling point**.

  - The **normal boiling point** is the boiling point of a liquid at 101.3 kPa.
  
    - Normal Boiling Point of Water - 100°C
Water Molecule

![Water Molecule Diagram]

- Polarity

![Polarity Diagram]

- Hydrogen Bonding

![Hydrogen Bonding Diagram]

Liquid State

- Surface Tension
  - Water molecules on the surface have unbalanced hydrogen bonding, and they are usually pulled inward. This inward pull is called surface tension.
  - Causes water droplets to form a spherical shape.
  - Surfactants
    - A substance that interferes with hydrogen bonding and reduces surface tension.
- **Vapor Pressure**
  - Because of strong hydrogen bonding, it is hard for molecules to escape from the surface of the liquid, therefore evaporation is slow and vapor pressure is very low.

**Water in the Solid State**

- **Density**
  - Usually, when a liquid gets cold, the molecules get closer together and it becomes more dense.
  - Ice is less dense than water because of hydrogen bonding. Crystals are formed with empty pockets of air.

- **Melting Point**
  - Because of strong hydrogen bonding in ice crystals, it takes a lot of energy to melt ice, so ice has a fairly high melting point.