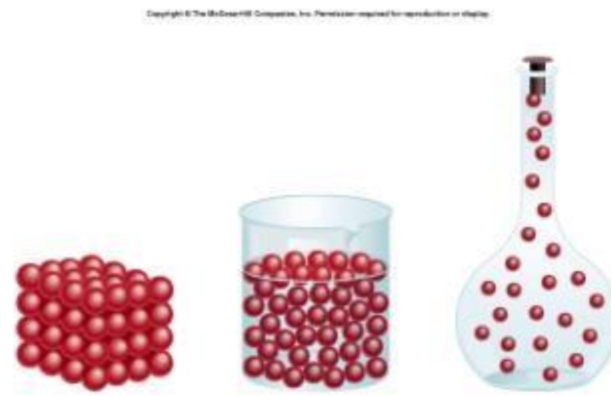
Kinetic Molecular Theory

Kinetic refers to things in motion; molecular deals with molecules; a theory is a hypothesis that has been supported with experimental evidence. Now, we need to define terms. You know what these terms are, but for the discussion of this unit, we will define them differently than normal.

Kinetic Molecular Theory is guided by the following assumptions:

1. The molecules of an ideal gas are in constant, random, straight-line motion. All matter is made of particles (atoms, ions, molecules). These particles are in constant motion because electrons are moving!
    a. Particles of a gas travel in completely *random* motion
    b. Particles of a liquid appear to vibrate around *moving* points
    c. Particles of a solid appear to vibrate around *fixed* points



2. The molecules in an ideal gas can be considered dimensionless points. This assumption can be made because molecules are very, very small compared with the empty space in a gas sample. Therefore, you can assume that the volume of a gas sample is equal to the size of the container that it is held.

    Example: If a student broke a bottle of cologne in the back of the lab and didn't tell anyone, you would be able to smell the cologne. The cologne's gas particles with spread out to fill the new container, which in this situation would be the classroom.

3. Collisions between molecules and the walls of its container are completely elastic. Elastic collisions mean that no energy is lost or gained when the molecules collide.

4. There are no attractive or repulsive forces in an ideal gas. This means that the molecules of an ideal gas do not like or hate each other, they are indifferent. If they pass by another gas molecule, they don't even notice them, they just go about their business. This assumption is totally false, but as long as the molecules of the gas have plenty of space, they really won't interact with each other.

Things that Affect Kinetic Energy of Molecules

The average kinetic energy of any molecules depends on their temperature. However, all molecules do not travel at the same speed, so even at one temperature, some molecules will be travelling slowly, while others will move more quickly. As the temperature is increased, the molecules will speed up and have larger kinetic energies. As a temperature decreases, the molecules will slow down and have smaller kinetic energies.

If an object could ever achieve absolute zero, all molecules would cease to move. Unfortunately, scientists have never reached absolute zero (-273K) in the lab.

Consequences of KMT (Kinetic Molecular Theory)

When gas molecules follow these rules, the gas is said to be "ideal". Most gases are ideal and will follow these rules. The only conditions under which these assumptions fail is when the gas sample is either under high pressure or at low temperatures. When a gas is compressed under high pressure, the gas is no longer mostly empty space, so assumption number 2 fails and the gas will condense to a liquid. When a gas is at low temperature, the molecules are moving slowly enough that they notice each other and assumption number 4 fails. Once the gas molecules realize that they are not alone, they start to feel attractive and/or repulsive forces and they will condense to a liquid.

Abbreviations, Conversions, and Standard Conditions:

Abbreviations

atm - atmosphere

mm Hg - millimeters of mercury

torr - another name for mm Hg

Pa - Pascal (kPa = kilo Pascal)

K - Kelvin

°C - degrees Celsius

Conversions

K = °C + 273

1 cm^3 (cubic centimeter) = 1 mL (milliliter)

1 dm^3 (cubic decimeter) = 1 L (liter) = 1000 mL

Standard Conditions

0.00 °C = 273 K

1.00 atm = 760.0 mm Hg = 101.325 kPa = 101,325 Pa

Vocabulary:

Temperature is a measure of the average kinetic energy of the molecules in a sample. Since $KE = \frac{1}{2} mv^2$, the change in temperature of a sample is caused by a change in velocity (speed) of the molecules in that sample.

Volume of a sample is the space that the particles take up (not much differently than it is normally thought of).

Pressure of a gas sample is caused by the molecules of that sample colliding with the walls of the container. Pressure = Force / Area, so in order for a gas to exert more pressure on its container, there must be more collisions or more forceful collisions.