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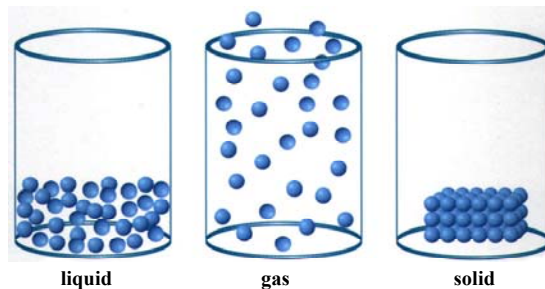
Chapter 10 The Kinetic Theory of Matter

10.1 Physical Behavior of Matter

A **solid** is rigid with a definite shape and volume, and its particles are fixed in place.

A **liquid** is flowing matter with a definite volume but an indefinite shape, and it takes the shape of the container.

A **gas** is flowing, compressible matter with neither definite volume or shape. Its particles are much farther apart and it takes the volume of the container.



The **Kinetic Theory** states that the tiny particles in all forms of matter are in constant motion.

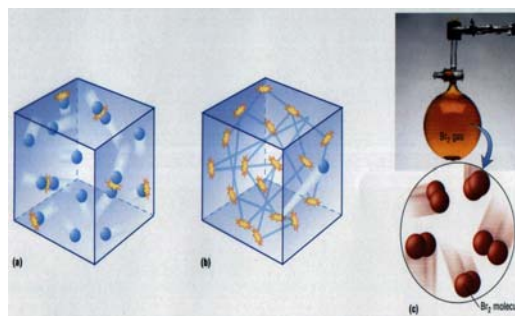
The word **kinetic** refers to **motion**

Kinetic energy (KE) is the energy an object has because of its motion

Kinetic Model of Ideal Gases:

1. Gases are composed of tiny particles (atoms or molecules), with insignificant volume. In-between is empty space without attractive forces between them.
2. Gas particles constantly travel fast in a straight line until they collide with one another or with the container wall.
3. All collisions are perfectly elastic, meaning the total kinetic energy remains the same.

Except at very low temperature and very high pressure, nearly all real gases behave like ideal gases



Gas pressure is a force exerted by a gas per unit surface area of an object.

Without (gas) particles \Rightarrow no collision \Rightarrow no pressure \Rightarrow **vacuum** (empty space)

Atmospheric Pressure results from collisions of air molecules with objects.

It decreases with elevation and also depends on the weather. At sea level, atmospheric pressure is about 14.7 psi = 14.7 pounds per square inch (about 1.75 gallons of milk pushing onto an area of 1 inch square).

Human and other forms of life have adapted to atmospheric pressure, but we are sensitive to pressure changes.

Kinetic Model of Liquids

Liquid particles are close together, so they can't move in straight lines anymore.

They are free to move past each other, but there are also attractive intermolecular forces between them. Particles of liquids also spin and vibrate while they move from place to place, which contributes to their kinetic energy.



Marbles: Model for liquids

Kinetic Model of Solids

Particles in solids cannot move past each other, but they vibrate in place. At the melting point particles start to move past each other.

Most solids are crystalline, with particles arranged in an orderly, repeating three-dimensional pattern called the **crystal lattice**.



Crystal lattice

Other Forms of Matter

Liquid Crystals When liquid crystals melt, they lose their rigid structure only in one or two dimensions, and stay arranged either in layers or parallel lines (used in LCD = liquid crystal displays).

Amorphous Materials (a-morphous = without structure) Amorphous solids have a disjointed and incomplete crystal lattice (e.g. wax, rubber, plastic, asphalt, glass)

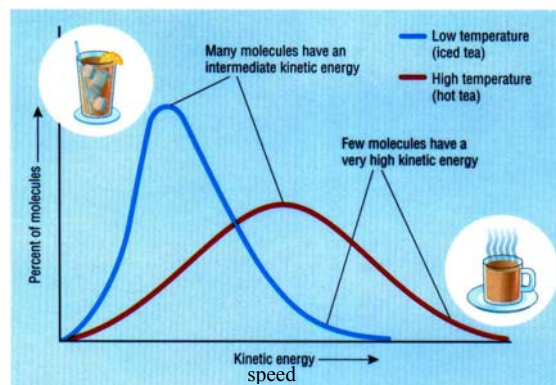
Plasma is ionized gas, but its electrically neutral because it contains equal amount of free electrons and positive ions. Most common form of matter in the Universe but least common on Earth, forms at very high temperatures and in fluorescent light tubes.

10.2 Kinetic Energy and Changes State

Average Kinetic Energy and Temperature

When a substance is heated, energy speeds up the particles, which increases the average kinetic energy of the substance, which increases the temperature.

Not all particles have the exact same speed or kinetic energy (distribution of KE similar to bell-shape curve of all test grades in a class).



Temperature of a material is a measure of the average kinetic energy

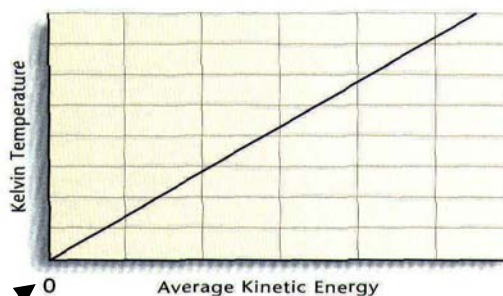
Higher Temperature = higher average KE = faster
Lower Temperature = lower average KE = slower

Kelvin Temperature Scale

Kelvin Temperature of a substance is directly related to the average kinetic energy of the substance (e.g. a substance has at 200K twice the average kinetic energy as it does at 100K) If you double the Kelvin temperature you double the average kinetic energy.

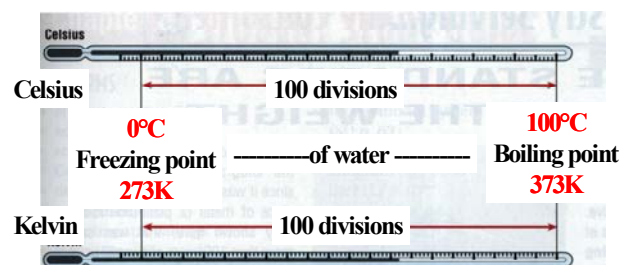
Absolute Zero:

When temperature is constantly decreased, eventually all particles should stop moving and have no kinetic energy = **Absolute Zero (zero Kelvin) = 0 K = -273.15 °C.**



Metric Unit for Temperature is Degree Celsius (reference points boiling point and freezing point of water)

SI Unit for Temperature is Kelvin (*no degree sign*) (reference points is absolute zero = 0 K = -273 °C)



Conversions: $K = ^\circ C + 273$ $^\circ C = K - 273$

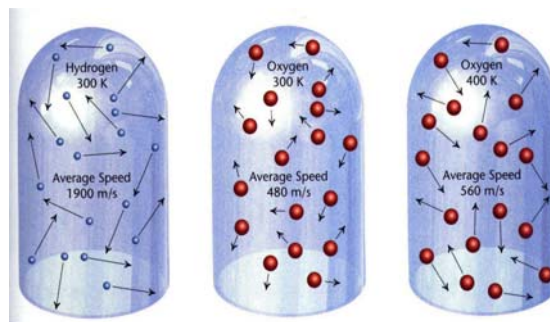
Convert 37°C to K

Convert 150 K to °C

Kinetic energy depends on speed and mass (compare the impact of being hit by a ping-pong ball or bowling ball, same speed).

$$KE = \frac{1}{2} m v^2 \quad (\text{kinetic energy} = \frac{1}{2} \text{mass} \times (\text{speed})^2)$$

At the same temperature the average kinetic energy for all samples is the same, but the speed will be lower if the mass of the particles is larger.



Diffusion

Diffusion is the process by which particles of matter fill a space because of random motion (example odor). The rate of diffusion depends upon the particles kinetic energy (mass and speed).

Changing State

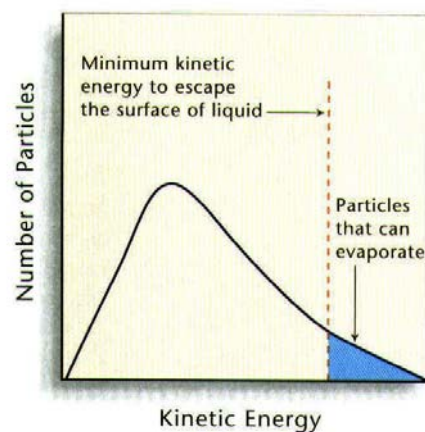
Vaporization: Conversion of liquid to gas (evaporation or boiling)



Evaporation: Conversion of particles from a liquid to a gas by escaping from the surface of a liquid below its boiling point. Only those molecules with enough kinetic energy can overcome the intermolecular forces and evaporate. This lowers average kinetic energy \Rightarrow cooling effect (e.g. perspiration).

Liquids that evaporate quickly are **volatile** liquids

Sublimation: Conversion of particles from a solid directly to a gas without being liquid in between (dry ice, moth balls, iodine)

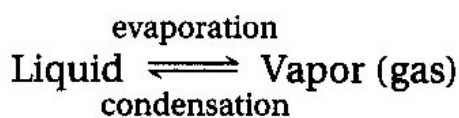
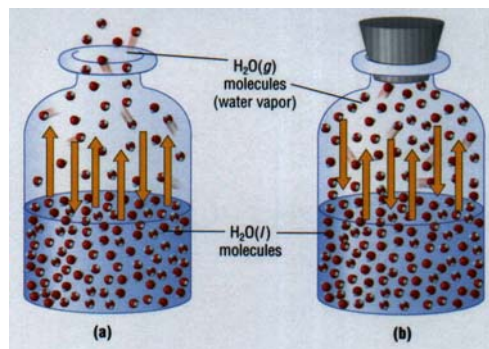


Condensation: Condensation is the reverse of evaporation. In condensation gaseous particles come close together (“*condense*”) and form a liquid.

Vapor Pressure:

Evaporated particles in a closed container collide with the container wall, producing a vapor pressure.

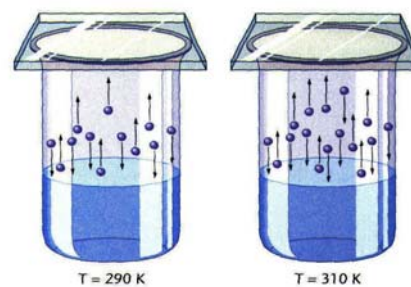
After a while, the number of evaporating particles will be equal to the number of condensing particles, called **equilibrium**.



Equilibrium

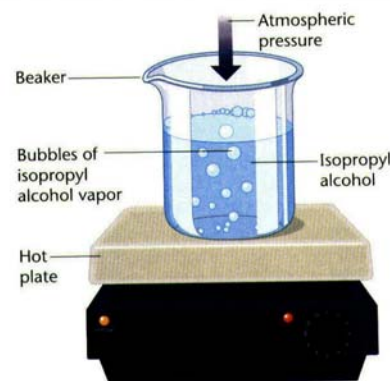
Increasing the temperature of a contained liquid increases its vapor pressure.

A liquid in an open container will never reach equilibrium, because the vapor is constantly escaping. But, exposed to the atmosphere, liquids having high vapor pressure will evaporate more quickly (more volatile) than those with low vapor pressure.



Boiling Point (bp): When a liquid is heated to a high enough temperature, many particles throughout the liquid have enough kinetic energy to vaporize. The boiling point is the temperature at which the vapor pressure of the liquid is equal to the external pressure. Bubbles of vapor form throughout the liquid rise to the surface and escape into the air. The boiling point of a substance changes with external pressure (lower pressure lowers the boiling point, because the particles need less kinetic energy to escape the liquid). Increasing the outside pressure raises the boiling point (ex. pressure cooker)

Normal boiling point is the boiling point at Standard atmospheric pressure, 101.3 kPa.



Heat of Vaporization

The energy absorbed when 1 kg of a liquid vaporizes at its normal boiling point is called its heat of vaporization. Because energy is always conserved, energy is released during the opposite process, when vapor changes back to a liquid. As the vapor condenses, the particles move closer together. You may know that burns from steam are usually more severe than burns from hot water. This occurs because the steam transfers a great deal of heat to the skin when it condenses (the same amount of heat as the heat of vaporization). Heat of vaporization is represented by the horizontal line in the figure below at 100°C.

Heat of Fusion

Heat of Fusion is the energy released when 1 kg of a liquid solidifies at its normal freezing point. Heat of fusion is represented by the horizontal line in the figure below at 0°C.

Changes in Temperature as a solid is heated

The figure on the right shows the rise in temperature of ice, water and steam (water vapor) as a sample of 100g is continuously heated. Notice the flat areas during changes of state when heat is supplied but the temperature does not change. This is, because all the energy is needed to overcome the intermolecular forces and break down the crystal lattice or convert a liquid into a gas.

