States of Matter

Chapter 13

Kinetic Theory

- **All matter consists of tiny particles that are always in constant motion.**

Kinetic Energy & Temperature

- **Temperature is a measure of the average kinetic energy of the particles.**
  - **Absolute zero: All motion stops**
  - **Zero Kelvin (0 K) = -273°C**
  - **Temperature in K = °C + 273**
  - **Standard temperature in "STP" = 0°C or 273K**

The Nature of Gases

- **Gas**
  - **No definite volume or shape**
  - **Particles of gas move rapidly (constant motion)**
  - **Elastic Collisions**

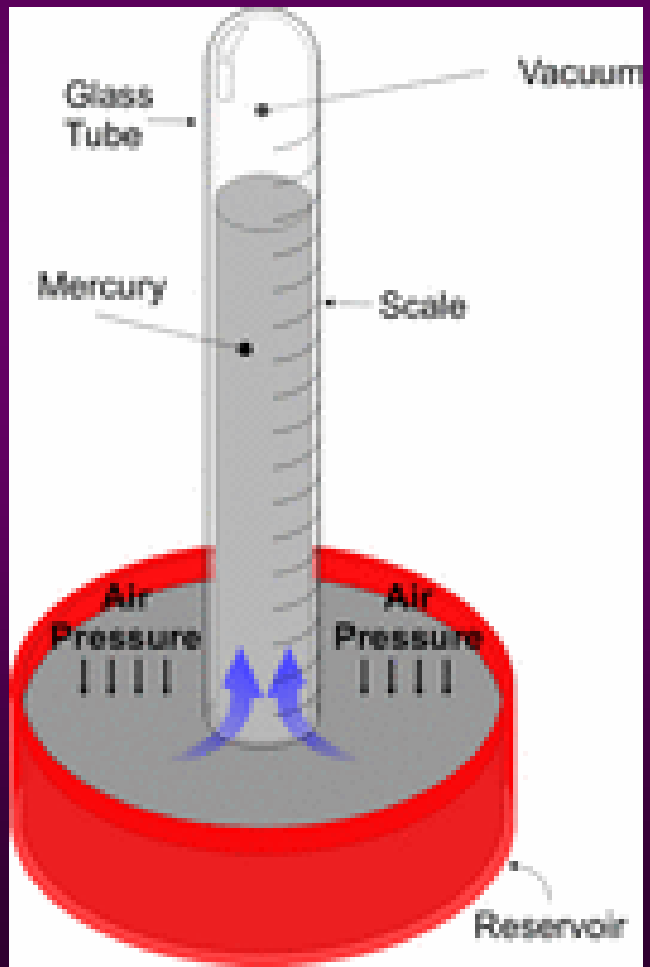
Gas Pressure

- **Result of simultaneous collisions of billions of rapidly moving particles in a gas with an object.**

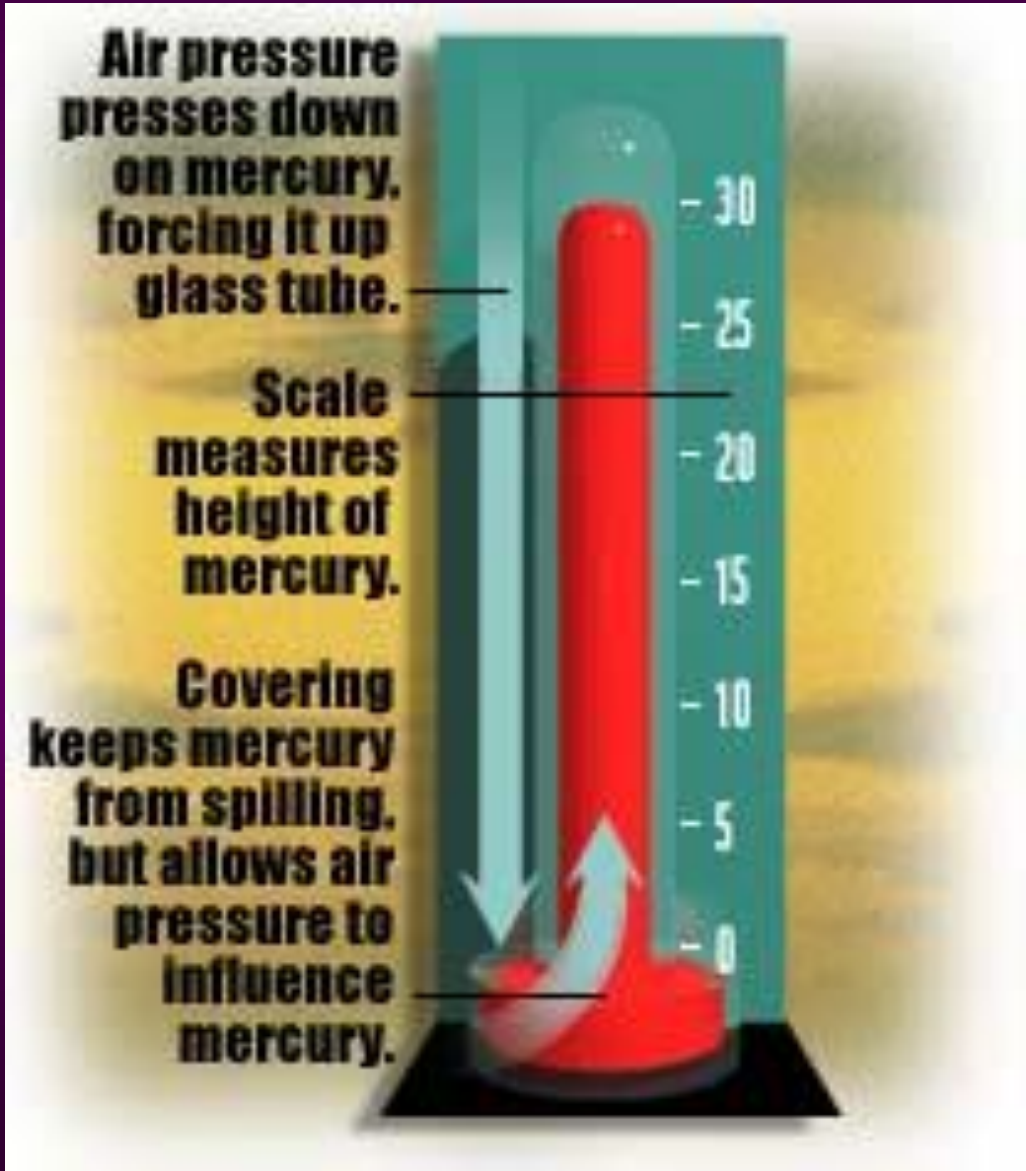
  - **<u>Atmospheric pressure</u>: results from atoms & molecules in the air colliding with objects.**

  - **<u>Barometer</u>: a device that measures atmospheric pressure.**

  - **<u>Vacuum</u>: An empty space with no particles present (no pressure).**



Barometer


Air pressure presses down on mercury, forcing it up glass tube.
Scale measures height of mercury.
Covering keeps mercury from spilling, but allows air pressure to influence mercury.
30
25
20
15
10
5
0

Nature of Gases

- **Standard Temperature and Pressure (STP)**
  - **Standard Temperature**
    - **0°C or 273K**
  - **Standard Pressure**
    - **1 atm**
    - **760 mm Hg**
    - **760 Torr**
    - **101.3 kPa**

Nature of Gases

- **Converting Between Units of Pressure**
  - **A gas is at a pressure of 1.50 atm**
    - **What is this in kPa?**
    - **In mm Hg?**
  - **Convert 190 mm Hg to:**
    - **kPa**
    - **atm**

Nature of Gases

- **Converting Between Units of Pressure**
  - **The pressure at the top of Mount Everest is 33.7 kPa. Is that pressure greater or less than 0.25 atm?**

Nature of Gases

- **Pressure & Elevation**

- **Atmospheric pressure decreases with increasing altitude (fewer particles available to collide)**

Nature of Liquids

- **Liquid**
  - **no definite shape; definite volume**
  - **Intermolecular forces keep the particles in a liquid close together**

Nature of Liquids

- **Evaporation or vaporization**
  - Conversion of liquid to gas
    - Particles at the surface of the liquid gain enough kinetic energy to overcome attractive forces between particles
    - Particles break away from the liquid and go into the gas or vapor state

Nature of Liquids

- **Vapor Pressure**
  - **In a sealed container, liquid particles evaporate into the gas phase above the liquid collide against the container**
  - **Dynamic equilibrium:**
    - **Rate of vaporization = Rate of condensation**

- **Boiling Point**
  - **Is the temperature at which the vapor pressure of the liquid = atmospheric pressure**
  - **Bubbles form throughout the liquid (“boiling”)**

Nature of Liquids

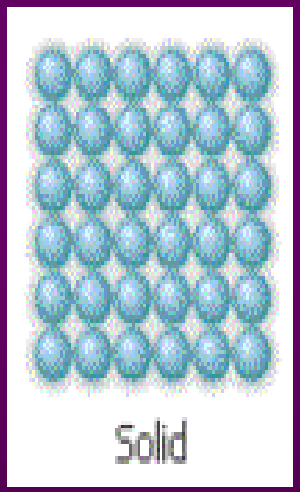
- **Normal Boiling Point**
  - **Is the boiling point of a liquid at standard pressure (101 kPa)**
  - **Water boils at 100°C at standard pressure**
  - **Water boils at lower temperatures when the pressure is lower (i.e. high altitudes)**
    - **Water boils at 95 °C in Denver (1600 m above sea level)**
    - **Takes longer to bake cake**
    - **Pressure cookers – increase pressure and BP of water!**

Boiling Point

- Does water always boil at 100ºC?
  - NO.
  - Boiling point depends on atmospheric pressure
    - Is the temperature at which vapor pressure of the liquid is the same as the atmospheric pressure
  - Pressure of the bubbles = air pressure in the room

Phase Changes

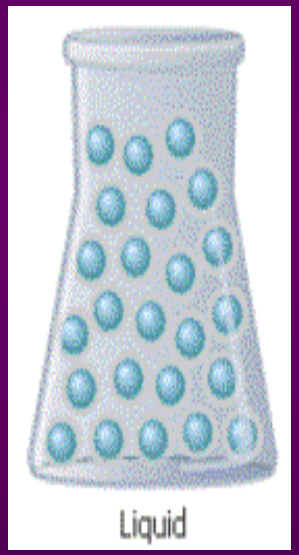
Sublimation



Melting

(fusion)

Freezing



Boiling

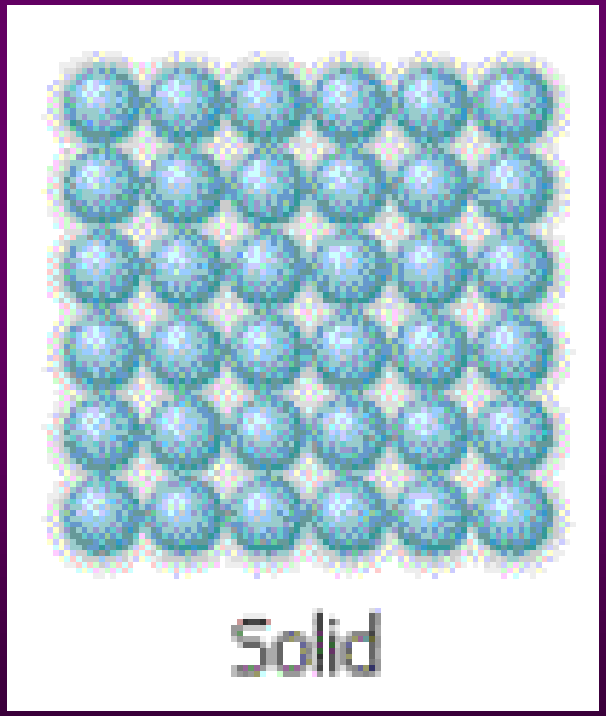
(evaporation or vaporization)

Condensation

Deposition

Nature of Solids

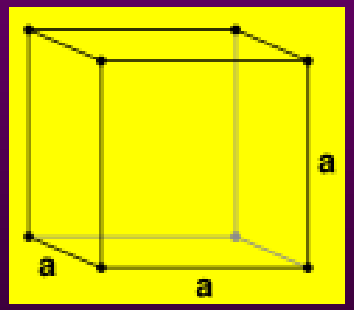
- **Solids**
  - **Definite volume and shape**



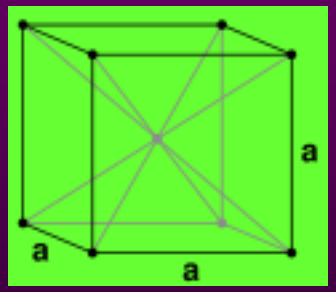
Nature of Solids

- **<u>Crystals</u> – atoms, ions or molecules arranged in orderly, repeating, 3D pattern**
  - **Unit cell: the smallest group of particles within a crystal that retains the geometric shape of the crystal**
  - **Examples:**

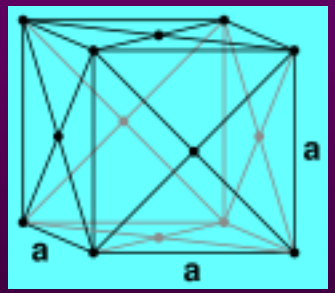
simple cubic



body-centered cubic



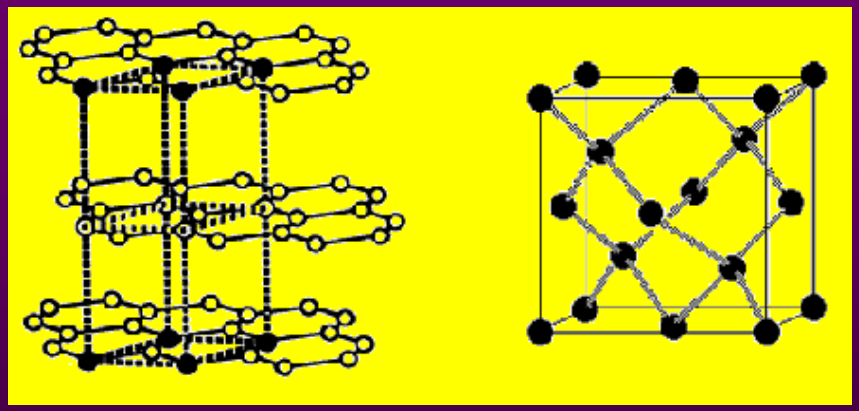
face-centered cubic



Nature of Solids

- **Allotropes – Two or more different molecular forms of the same element in the same physical state**
  - **Example: carbon**

graphite **diamond**



- **Amorphous solids – lack ordered structure**
  - **Examples: glass, plastics, rubber**

Nature of Solids

- **Melting Point – temperature at which a solid becomes a liquid**

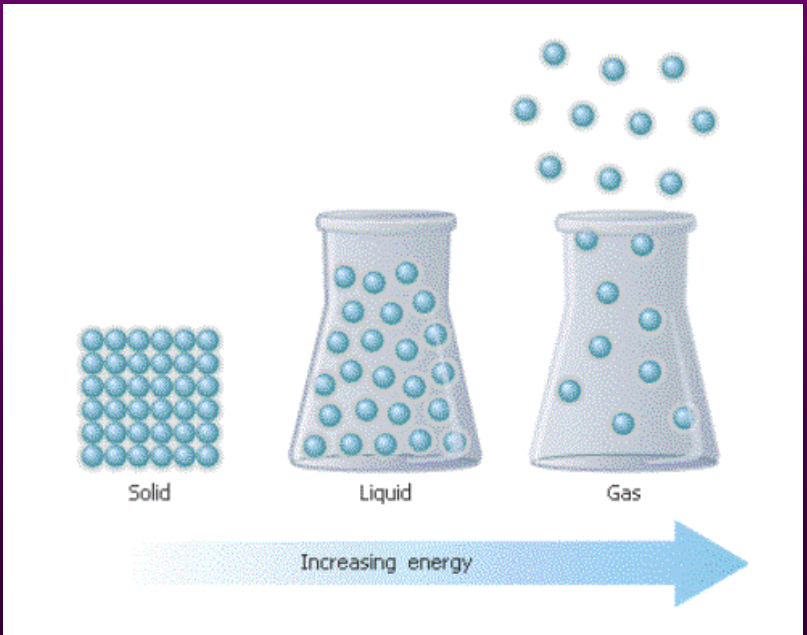
- **Freezing Point – temperature at which a liquid becomes a solid**

Sublimation of Iodine

- **Sublimation**
  - **Solid changes to gas without going through the liquid phase**
- **Demonstration**
  - **Solid iodine changes to a gas when heated**

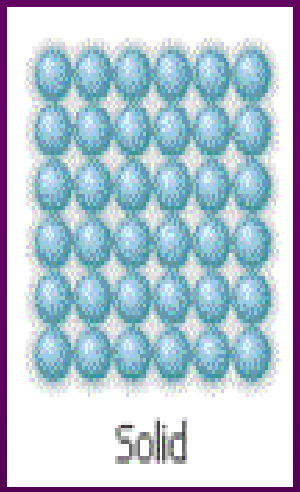
Phases of Matter

- From a solid to a liquid:
  - Temperature increases (motion of particles increases)
  - Attraction between particles decrease



Phase Changes

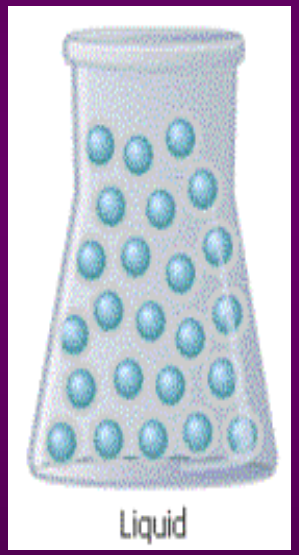
Sublimation



Melting

(fusion)

Freezing



Boiling

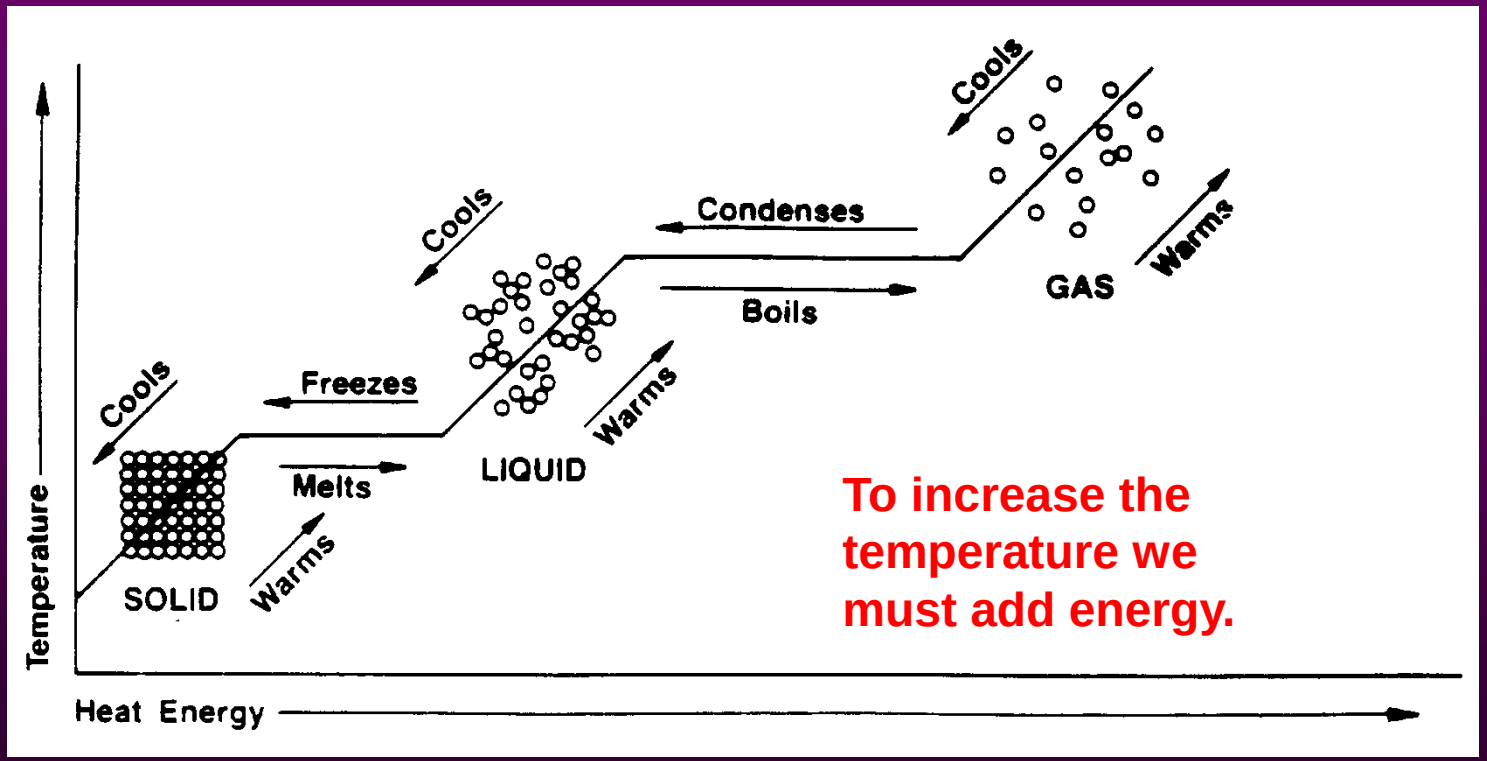
(evaporation or vaporization)

Condensation

Deposition

Heating/Cooling Curve

- **Shows the temperature and energy of a substance over time as it changes from a solid to a gas**



Phase Diagram of Water

- **Shows the relationship among pressure, temperature, and the 3 phases of water (gas, liquid & solid).**
- **Triple Point – conditions at which all 3 phases can exist in equilibrium together**