CHAPTER 13

States of Matter

Kinetic Molecular Theory

States that the tiny particles in all forms of matter are in constant motion.

Kinetic = motion

13.1 THE NATURE OF GASES

- A gas is composed of particles, usually molecules or atoms, with negligible volume *(Volume not worth considering)*
- Gases have a small or no volume.
- Gas particles are far away from each other and are separated by large amounts of empty space between them.

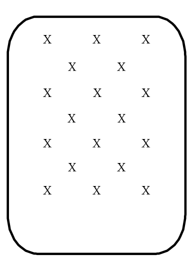
Polar vs. Nonpolar

Polar Molecules
Have an attraction to each other
Stay closely packed
Liquids

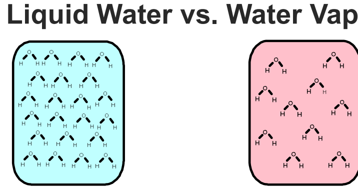
NonPolar Molecules
Have no attraction to each other
Stay Spread out
Gases

Gases

- H_2, O_2, CO_2, N_2
- Gases are Non-Polar Molecules
- Molecules are NOT attracted to each other and have empty space between them.
- Allows them to take the shape of their container and flow.



Liquid Water vs. Water Vapor



Low Kinetic Energy
Polar molecules highly attracted

High Kinetic Energy
Polar molecules not attracted

Gas Particle Motion

- The particles in a gas move randomly and rapidly in a straight path until collisions occur with each other or objects.
- Aimless path they take is called a random walk.
- All collisions are perfectly elastic

- When elastic collisions occur, particles bounce and energy is transferred from one particle to another, but the total kinetic energy remains constant.

Particle Energy

- Kinetic Energy is determined by mass and velocity of particles

$$K = \frac{1}{2}mv^2 \qquad K.E = \frac{1}{2} m v^2$$

- Velocity represents speed and direction.
- All particles do not have the same KE since their molar masses may be different.

Calculating Kinetic Energy

A gas particle of mass 3.25×10^{-24} kg has a velocity of 1.00×10^3 m/s.

$$K = \frac{1}{2}mv^2$$

What is the kinetic energy of the molecule?

1.63×10^{-18}

What temperature scale is best when discussing KE?

Kelvin

KINETIC ENERGY AND TEMPERATURE

An increase in temperature causes the average kinetic energy of all gases to increase.

- Kelvin Scale and Kinetic Energy are proportional

Absolute zero is the temperature at which the motion of particles theoretically ceases or stops.

- **0 K = -273º C**
- **0 Kelvin means NO kinetic energy**

Temperature

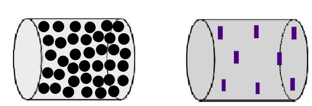
- A measure of the average kinetic energy that particles possess in a sample of matter.
- We use average because different gases have different molar masses, hence, different KE. For example, you have a sample of gas consisting of different types of gas molecules.
- How could gases with different molar masses have the same KE and temperature? They have a different Velocity

Behavior of Gases

Gases take the shape of their container

Gases have low density due to the great deal of space between particles. This means there are less particles per a given volume

1-L Gold @ 20°C



1-L chlorine @ 20°C

Compression vs. Expansion

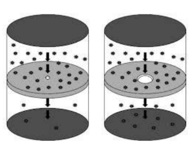
Reducing the volume of a particular mass of gas, resulting in less empty space between particles.	Increasing the volume of a particular mass of gas, resulting in more empty space between particles.
Increases density	Decreases density

Diffusion

- Gases mix until they are evenly distributed.
- Gases move from areas of high concentration to areas of low concentration.
- Gases with low molar mass diffuse fastest.
- Aerosol cans, oxygen tanks, air compressors.

Effusion

- Gases escape through *tiny holes* in it's container.
- Gases with low molar mass effuse the fastest.
- Balloons, tires, blimps, floats



- N_2 = 28-g/mol (78%)
- O_2 = 32-g/mol (21%)
- Ar = 40-g/mol (0.934%)
- CO_2 = 44-g/mol (0.033%)
- Ne = 20-g/mol (0.0018%)
- He = 4-g/mol (0.00052%)

Graham's Law of Effusion

The rate of effusion for a gas is inversely proportional to the square root of it's molar mass

$$\frac{\text{Rate of A}}{\text{Rate of B}} = \sqrt{\frac{\text{mass of B}}{\text{mass of A}}}$$

$$\text{Rate}_{\text{effusion}} \propto \frac{1}{\sqrt{\text{density}}} \propto \frac{1}{\sqrt{\text{MM}}}$$

Graham's Law of Effusion (Example 1)

Compare the effusion rates for Bromine and Krypton at the same temperature and pressure.

$$\frac{\text{Rate Kr}}{\text{Rate } Br_2} = \sqrt{\frac{\text{Molar Mass } Br_2}{\text{Molar Mass Kr}}} = \sqrt{\frac{160\text{ - g/mol } Br_2}{84\text{ - g/mol Kr}}} = \sqrt{1.90} = \boxed{1.38}$$

Kr effused 1.38 times faster than Br_2

If it took 40 seconds for the bromine to effuse, how long did it take for the krypton to effuse?

Graham's Law of Effusion (Example 2)

An unknown gas effuses 4.0 times faster than O_2. What is the molar mass?

$$\frac{\text{Rate A}}{\text{Rate B}} = \sqrt{\frac{\text{Molar Mass B}}{\text{Molar Mass A}}}$$

$$\frac{\text{Rate A}}{\text{Rate B}} = 4.0$$

$$4.0 = \sqrt{\frac{32\text{ - g/mol } O_2}{M_{\text{unknown}}}}$$

Square both sides cancels $\sqrt{\ }$

$$16.0 = \frac{32\text{ - g/mol } O_2}{M_{\text{unknown}}}$$

Cross multiply

$$\frac{32\text{ - g/mol } O_2}{16} = 2\text{ - g/mol}$$

Graham's Law of Effusion (Example 4)

Which of the following gases effuses 3.74 times slower than hydrogen? (CO_2, NH_3, O_2, N_2, C_2H_6)

$$\frac{\text{Rate A}}{\text{Rate B}} = \sqrt{\frac{\text{Molar Mass B}}{\text{Molar Mass A}}}$$

Gas Pressure

- Force exerted on an object by a gas.
  1. Heat causes particles to move master = more collisions.
  2. Number of particles increases = more collisions
- The result of simultaneous collisions of billions upon billions of gas particles with an object.
- Hair Spray Can, Balloons
- The gas inside exerts pressure on the walls of the container

Vacuum

- An empty space without matter.
- Has NO gas molecules.
  > No molecules = no collisions
- Outer space has no air pressure
- Pressure = 0.0 kPa

Atmospheric Pressure

- Pressure of our atmosphere
- Results from the collisions of air molecules with objects.
- Barometers are used to measure atmospheric pressure.
- Manometers measure pressure in a closed container
- SI unit of pressure is the pascal (Pa)

Atmospheric Pressure

- As you increase altitude in our atmosphere, the atmospheric pressure decreases
- Less force is exerted by the air particles.
- Due to gravity, there is less air at high altitudes. *(Thin air)*
- Density of Earth's atmosphere decreases as altitude increases.

Altitude Changes

- What happens when you change altitudes with a half-filled water bottle with the cap sealed tight?

- Increase altitude from sea level, the the bottle expands.
- Decrease altitude from the top of a mountain, the bottle collapses.

Units of Pressure

- 101.3 kPa
- 1 atm
- 760 mm Hg
- 760 Torr

Same amount
Different units

STP: standard conditions for gases
0ºC and 101.3 kPa of pressure

Converting between units

- Convert 375 torr into atmospheres
- Convert 989 mm Hg into kilopascals

Dalton's Law of Partial Pressures

- used for a mixture of gases
- The contribution each gas in a mixture makes to the total pressure
- The pressure each gas contributes to the total pressure is the partial pressure.

Dalton's Law of Partial Pressures *(cont.)*

$$P_{total} = P_1 + P_2 + P_3$$

Find the total pressure of a mixture of 3 gases with partial pressures of 10.0-kPa, 16.7-kPa, and 34.9-kPa.

O_2 CO_2 N_2

Dalton's Law of Partial Pressures *(cont.)*

A mixture of O_2, N_2, He, & CO_2 have a total pressure of 1.07- atm. What is the partial pressure of CO_2 if the $P_{nitrogen}$= 0.62, P_{oxygen}=0.29, and P_{helium}=0.04.

13.2 Forces of Attraction

- Intermolecular Forces: the attraction between particles.
- Holds molecules to each other.
- Weaker than ionic or covalent bonds.
- They are responsible for whether a molecular compound is a gas, liquid, or solid.

1. Dispersion Forces
2. Dipole - Dipole Forces
3. Hydrogen Bonds

Gases	Liquids	Solids
No attraction	Dipole Attraction	Ionic Attraction
Nonpolar Molecules	Polar Molecules	Ions Form Crystals

Van der Waals forces

- The weakest attractions between molecules. **Not Bonds!!!!!!**
- Three types:
  > Dispersion forces, Dipole interations, and Hydrogen bonds
- Hydrogen > Dipole > Dispersion *(Strongest to weakest)*

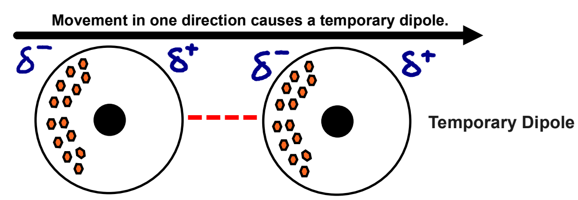
Attractions between polarized molecules

Dispersion Forces

- The weakest of all intermolecular interactions.
- Thought to be caused by the motion of electrons.
- Strength of dispersion forces increases as the number of electrons in a molecule increases
- Electrons are not lost or gained

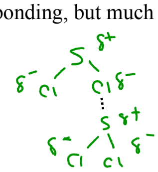
Dispersion forces

Due to movement, the electrons move to one side and create a separation of charge.



Dipole Interactions

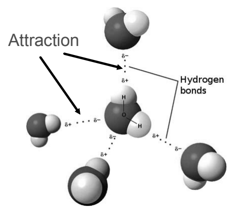
- Occur when polar molecules are attracted to one another
- Electrostatic attractions occur between theoppositely charged regions of dipolar molecules.
- Similar to ionic bonding, but much weaker attraction.



Hydrogen Bonds

- Dipole interactions involving hydrogen.
- Strongest intermolecular attraction.
- The covalently bonded hydrogen becomes slightly positive.
- Unshared electron pairs and atoms with high electronegativity become attracted to the slightly(+) Hydrogen.
- An atom or molecule is attracted to a Hydrogen atom that is already bonded to an atom with high electronegativity.

Hydrogen Bonding in Water



Hydrogen Bonding is the attraction between polar molecules with hydrogen.

Why is there so much water?

Water molecules are polar.

The oxygen atom becomes slightly negative and each hydrogen becomes slightly positive.

Causes an intermolecular attraction between water molecules.

The attraction water molecules have for one another is called Hydrogen bonding.

13.3 LIQUIDS

- Liquids flow = fluidity
- Take shape of container, but do not expand
- Held together by intermolecular forces between polar molecules, and have no empty space between them
- Most particles do not have the energy to overcome the intermolecular forces and escape into the gas phase
- No bonding between particles

Density & Compressibility

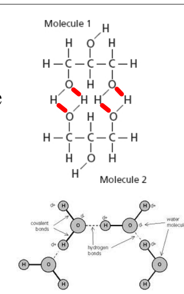
- If the temperature and pressure of a liquid and gas are the same, the density of the liquid is far greater than that of a gas.
- Due to intermolecular attractions...
- Liquids can be compressed, but not easily.

Fluidity of Liquids

- Liquids and gases flow and take the shape of their container.
- Liquids flow slower than gases at the same temperature due to their intermolecular attractions.
- Increasing temperature allows liquids and gases to diffuse and flow faster

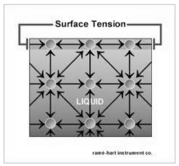
Viscosity

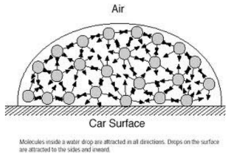
- Measure a liquids resistance to flow
- High viscosity liquids stronger attractive forces. More H-bonds, more viscosity.
- Large molecules with long chains have the highest viscosity due to the greater likelihood of attractions between atoms.
- Temperature increases lower viscosity!



Surface Tension

- The inward pull from particles on the inside. The stronger the attraction between particles, the greater the surface tension.
- Compounds that lower surface tension are surfactants.

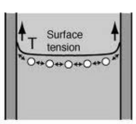


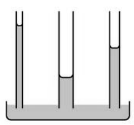


- The surface tension of water may be decreased by adding a wetting agent such as soap or detergent.
- Soaps or detergents are **surfactants.**
- Surfactants interfere with the hydrogen bonding between water molecules and stop them from beading.

Capillary Action

- When water is placed in a narrow container.
- Water becomes more attracted to glass than other water molecules, therefore, climbing the walls of the tube.





SOLIDS

- Fixed position = do not flow
- STRONG Attraction between particles
- Most are crystals

Density of Solids

Particles in a solid are more closely packed than liquids.

Most solids are more dense than most liquids

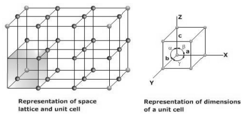
- Most particles become more dense when they freeze, except water does the opposite due to H-bonding.

Crystalline Solids

- Most solids are crystals
- In a crystal, the atoms, ions, or molecules are arranged in an orderly, repeating, three dimensional pattern called a crystal lattice.
- 6 types of Solids

Unit Cell

- The smallest group of particles within a crystal that retains the geometric shape of the crystal is known as a unit cell.
- The shape of a crystal depends on the arrangement of the particles within it. **A cube is the unit cell, not the particle.**
- 12 particles/cube, 100 cubes/cube crystal



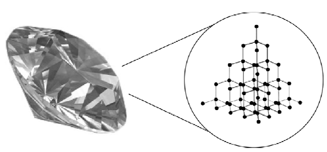
1. Molecular solids

- Molecules are held together by intermolecular forces and therefore have low melting points (Butter)
- Most are not solid at room temperature.
  > (Saturated vs unsaturated fats)
- Sugar is solid due to large molar mass
- Not all solids melt, some decompose (wood)

2. Network solids

Atoms are held together by many covalent bonds.

High Melting points and very hard



Allotropes

- Some substances can exist in more than one type of solid state
- Allotropes are two or more different molecular forms of the same element in the same physical state.
- Carbon can be a Diamond or Graphite
- Both are made of carbon and are solids, but diamonds are hard and graphite is soft.

3. Ionic solids

- Held together by oppositely charged ions
- In general ionic solids have high melting points because these solids are held together by strong forces.
- Tend to be hard and brittle.

4. Metallic solids

- Positive metal ions surrounded by mobile electrons
- Excellent conductor of electricity

5. Atomic solids

- Soft an have low melting points
- Atomic solids are formed by the noble gases

6. Amorphous solids

- Lack an ordered internal structure.
- Can change shape based on surrounding pressure
- Glasses are amorphous solids that are sometimes called supercooled liquids.
- Not like Crystal solids.

13.4 Phase Changes

- Some phase changes require energy
  > Melting and vaporization, sublimation
- some phase changes occur when energy is released.
  > Condensation, Freezing, Deposition

Melting Point

- The point at which a solid melts and becomes a liquid.
- Requires a gain in kinetic energy.
- Occurs at the same point at which a liquid freezes to become a solid.

EVAPORATION

- Evaporation (vaporization) is the conversion of a liquid to a gas or vapor.
- Occurs at the surface of a liquid.
- Liquid does not need to be boiling!
- The particles with the most kinetic energy evaporate first

EVAPORATION

- Evaporation is a cooling process
- Heating increases vaporization
- The liquid on the surface is exposed to the greatest amount of heat, so they are most likely to evaporate
- **Makes the Temperature of the liquid decrease!!!**

BOILING POINT

- The boiling point is the temperature at which the vapor pressure of the liquid is just equal to the external pressure.
- Water begins to boil.
- If external pressure increases, then boiling point increases.
- If external pressure decreases, then boiling point decreases.

External Air Pressure
93.7 - kPa

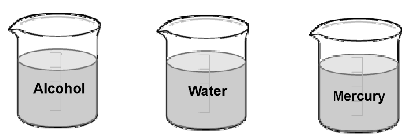
Vapor Pressure

- Pressure that a liquid boils
- All liquids at a given external pressure have the same vapor pressure.

Boiling Point

- Temperature at which a liquid reaches it's vapor pressure
- If liquids have the same vapor pressure, it doesn't mean they have the same boiling points

External Air Pressure 150.0 - kPa



1. What is the vapor pressure of all 3 liquids? ________________

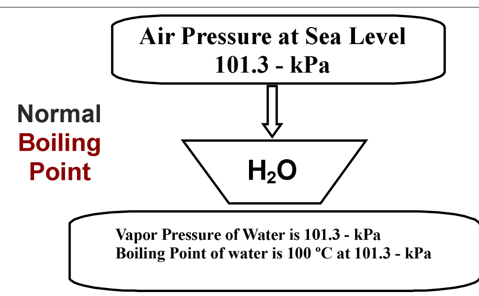
2. Alcohol evaporates fastest, which means it has the lowest ________________.

How can you raise the boiling point of water?

- Raise the external pressure
- Sanitization: water boiling at a higher temperature will sanitize things better.

Normal Boiling Point

- Boiling point of a liquid at sea level:
  - 101.3 kPa
  - 1 atm
  - 760 Torr
- Unique for each substance.
- The temperature of a boiling liquid never rises above its boiling point!!! Why?



Air Pressure at Denver
84.3 - kPa

⇩

H_2O

Vapor Pressure of Water is 84.3 - kPa
Boiling Point of water is 95 °C at 84.3 - kPa

Air Pressure in a Vacuum
0.0 - kPa

⇩

H_2O

Vapor Pressure of Water is 0.0 - kPa
Boiling Point of water is 0 °C at 0.0 - kPa

Condensation

- Water vapor molecules lose energy and their velocity is reduced, allowing H-Bonds to reform.
- Liquid is then produced.
- Occurs when water vapor meets cold air or cold objects

Dynamic Equilibrium

- Takes place in a closed container in which the vapor pressure remains constant!!
- As some liquid molecules evaporate, other gas molecules condense to maintain constant pressure.
- Rate of evaporation = Rate of condensation
- Each sample of liquid will need to reach a certain pressure before it is at Dynamic Equilibrium.

Freezing

- When liquids become a crystalline solid
- Occurs when energy has been removed from a liquid and velocity of particles decreases.
- Molecules have a fixed position.

Sublimation:The change of a substance from a solid to a gas or vapor without passing through the liquid state.
- This process requires energy.
- Iodine, Dry Ice

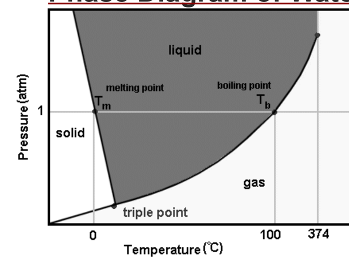
Deposition:The change of a substance from a gas or vapor to a solid without passing through the liquid state.
- This process releases energy
- Rapid crystallization of water vapor high in atmosphere

PHASE DIAGRAMS

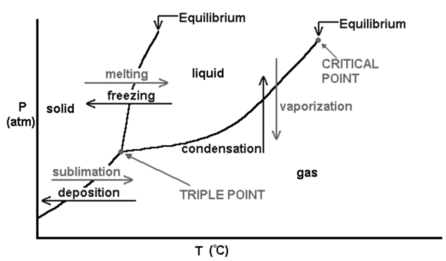
The relationships among the solid, liquid, and vapor states or phases of a substance in a sealed container are best represented in a single graph called a phase diagram

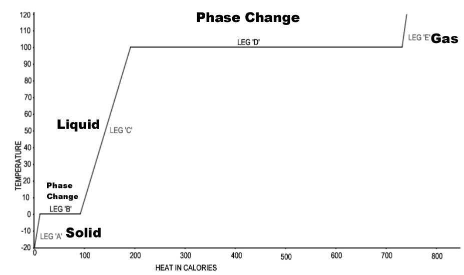
More energy is required for a phase change than a temperature change!!!

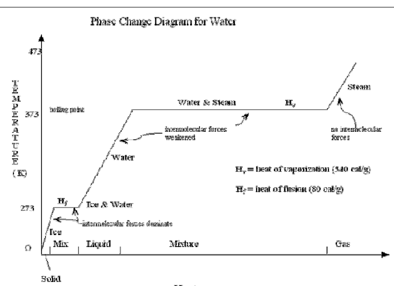
Phase Diagram of Water



Phase Diagram Explanation







Triple Point

- the point at which all three phases can exist in equilibrium with one another
- There is only 1 triple point for each substance