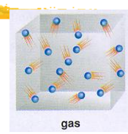
Chapter 11

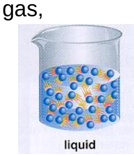
- KMT for Solids and Liquids
- Intermolecular Forces
- Viscosity & Surface Tension
- Phase Changes
- Vapor Pressure
- Phase Diagrams
- Solid Structure

Kinetic Molecular Theory

Liquids and solids will experience some of these same properties as a gas, with differing results:

- compressibility
- density
- shape & volume
- diffusion
- expansion during heating





Attractive Forces

When we discuss the state of matter, we must discuss the strength of attractions between its particles. What is this relationship between the particles attraction and their KE?

The lower the KE of the particles, the more the attractive forces hold the matter together.

Intramolecular Forces

Most liquids and solids are held together by covalent bonding.

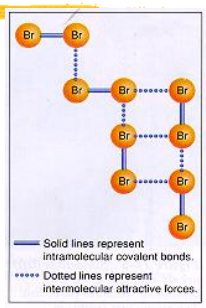
Covalent bonds hold the individual atoms of the molecule together. This is termed as an intramolecular force.

(intra - inside or within)

Intermolecular Forces

The forces that hold the molecules together in a liquid or a solid are called intermolecular forces. These are the forces of attraction between neighboring molecules. (inter - outside)

Which is stronger, inter-molecular or intra-molecular forces?



Forces that hold solids & liquids together

Intermolecular Forces

- Dispersion (London) Forces
- Dipole-Dipole Forces
- Hydrogen Bonding
- Ion-Dipole Forces

Bonds

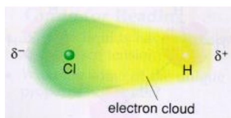
- Ionic Bonds
- Metallic Bonds
- Covalent Network Bonds

Dipole-Dipole Forces

Occurs inside a polar covalent molecule.

Due to differences in electronegativities, atoms within a molecule develop a partial charge.

The partial negative charge will attract the partial positive charge of another molecule, holding them together

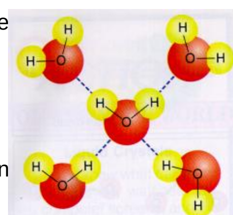


Hydrogen Bonding

Special type of dipole-dipole forces

When bonds with H and the 3 most electro-negative nonmetals occur, a very large partial charge occurs.

These bonds are very polar, much more so than dipole-dipole forces, so they are given a special name.

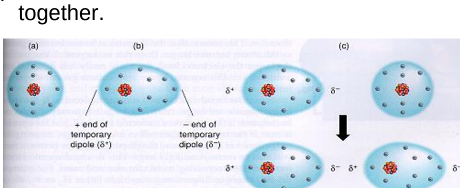


London Dispersion Forces

Starts inside an atom or nonpolar molecule

Atoms become temporary dipoles when more electrons are on one side of the atom than the other.

These temporary dipoles attract other temporary dipoles, and that attractive force holds them together.



London Dispersion Forces

Dispersion forces are present in all solid and liquid materials.

Dispersion forces are stronger with increasing atomic mass.

Ion-Dipole Forces

Special kind of IF for mixtures:

When you have a mixture of polar and ionic compounds, like in water solutions, an ion is attracted to the opposing charges of the dipole (water). The strength of the attraction increases with a greater charge.

Example: NaCl (aq)

Effects of Intermolecular Forces

Viscosity:

The “friction” or resistance to movement of the molecules as they move past one another.

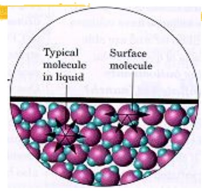
The more viscous a liquid, the stronger the intermolecular forces are that hold the liquid together.

Viscosity increases as temperature decreases.

Effects of Intermolecular Forces

Surface Tension:

Molecules at the surface of a liquid experience an uneven force of attraction, which causes a tight film that is stretched across the surface.



Surface tension also increases as the temperature goes down.

Vapor Pressure

Some liquids readily evaporate into a gas. This gas exerts a pressure on its surroundings. We call this pressure from a volatile liquid the vapor pressure.

Molecules within a liquid move at various speeds, and a fast moving particle within the liquid could have enough kinetic energy to break through the surface tension and leave the liquid (evaporate)

Vapor Pressure & Boiling

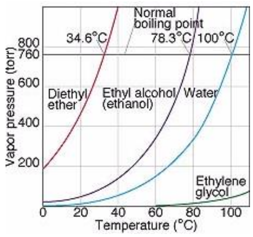
Vapor pressure increases as the temperature rises, due to faster molecule movement.

When the vapor pressure (pressure of gas molecules escaped from a liquid) is equal to the atmospheric pressure (pressure holding the liquid down), a liquid will begin to boil.

The normal boiling point is the temperature the liquid needs to be to have a vapor pressure equal to 1 atm.

Vapor pressure and Boiling

If the atmospheric pressure is different than 1 atm, the boiling point of the liquid will change.



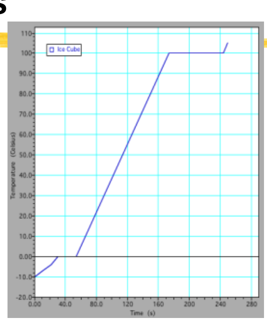
Heat of Vaporization & Fusion

The temperature of the substance does not increase until all of the solid (or liquid) has melted (or boiled off). The amount of heat needed to melt and boil a substance is called its heat of fusion and heat of vaporization.

ΔH_f = 334 J/g for water

ΔH_v = 2440 J/g for water

Heating Curves



As something is heated from a solid to a liquid to a gas, there are two abnormalities to the curve.

The two abnormalities occur at the fusion and vaporization (melting and boiling) points of the substance.

Overcoming and Succumbing to Intermolecular Forces

Phase changes:

Solid to Liquid ---> Melting

Liquid to Solid ---> Freezing

Liquid to Gas ---> Vaporization

Gas to Liquid ---> Condensation

Solid to Gas ---> Sublimation

Gas to Solid ---> Deposition

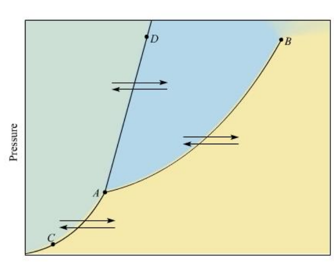
Phase Diagrams

A graphical way to summarize the pressure and temperature conditions under which equilibria exist between different states.

The triple point is the intersection of the three lines - where all three states exist in equilibria

The critical point - pressure and temperature limits of a substance - above these points it is impossible to tell if a substance is a liquid or a gas.

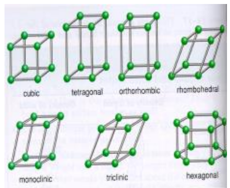
Phase Diagram



Solids

Crystalline solids - A solid in which the representative particles are in a very ordered, repeating pattern.

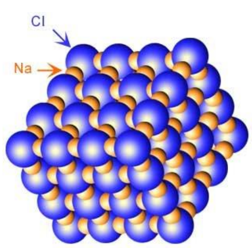
Amorphous Solids - A solid in which there is no orderly structure. Intermolecular forces vary in strength through the sample and soften over a large temp. range.



Bonding in solids

Molecular solids - atoms or molecules held by weak intermolecular forces, low to medium melting points. (Ar, H_2O, CO_2)

Ionic Solids - ions held together by ionic bonds, high melting points (NaCl, CaF_2)



Sodium Chloride Crystal Structure

Bonding in solids

Covalent-network solids - atoms held together by a network of covalent bonds, very high melting points. (graphite, diamond)

Metallic Solids - Metal atoms held together by metallic bonds, positive metal nuclei surrounded by a “sea of delocalized electrons”. Low to very high melting points. (Al, Fe, Cu, Ag, Au)