LIQUIDS / SOLIDS / IMFs

**Intermolecular Forces (IMFs)**

- **NOT** within molecules
- **NOT** true bonds... weaker attractions
- Represented by dashed lines
- Physical properties (melting points, boiling points, vapor pressures, etc.) can be attributed to the amount and strength of IMFs

Example: **HIGHER BP = STRONGER IMFs**

**InTRaMolecular Force**

- TRUE BONDS that hold atoms together... STRONGER!

**InTERMolecular Force**

- attractive force that keeps molecules together... WEAKER!

**London Dispersion Forces**

- **INSTANTANEOUS** dipoles due to random movement of e- (because of collisions with other molecules)
- **Exist in all molecules even NONPOLAR**
- Creates momentary, short-lived nonsymmetrical distribution of charge (dipole) in neighboring molecule

**Dipole-Dipole Forces**

- Attractions between molecules with dipoles (POLAR)
- Molecules arrange themselves so opposite poles align
- Larger the dipole, stronger the force of attraction
- **LITTLE STRONGER**... only 1% as strong as ionic bonds
Dipole-Dipole Forces

Attractive forces **DISSIPATE** with **INCREASED** distances!

Ex: Gases

Hydrogen Bonding

- Special type of dipole-dipole attraction... Hydrogen bonded to N, O, or F
- Attraction strength is higher because H bonded to highly electronegative elements
- Opposite poles align
- **STRONGEST** of the IMFs
- Important for bonding and properties of water and DNA

Hydrogen Bonding

\[ \delta^- \quad H \quad \delta^+ \]

Hydrogen bond

\[ \text{Notice that } \text{H}_2\text{O, HF, and } \text{NH}_3 \text{ are HIGHER than they should be!} \]

- Expect that BP would increase with increasing molar mass (more polarizable cloud)... Stronger IMFs mean higher BP
- BUT lighter hydrides have H bonding and stronger IMFs and small size which means closer, stronger dipoles!

Hydrogen Bonding

Liquid State Properties

- **SURFACE TENSION**: resistance of a liquid to increase its surface area / beading of liquid droplets

**HIGHER IMFs = GREATER!**

Intermolecular Forces

- **EXAMPLE:** Identify the main type of IMF in each.

  - CCl$_4$
  - CH$_3$OH
  - C$_4$H$_{10}$
  - HF
Liquid State Properties
• **CAPILLARY ACTION**: spontaneous rising of a liquid in a narrow tube
  
  - **COHESIVE FORCES**: between molecules
  - **ADHESIVE FORCES**: between molecules and container
  - Water... Adhesive **GREATER** than cohesive (concave meniscus)... Hg would be convex!

Liquid State Properties
• **VISCOSITY**: resistance to flow / increases with greater size and lower temps too

Liquid State Properties
• **VOLATILITY**: ability to evaporate readily

Solution Formation and IMFs
• In order to dissolve a substance in a solvent:
  1) Must overcome *(requires energy)*
     - Solute-solute IMFs
     - Solvent-solvent IMFs
  2) Form solute-solvent attractive forces upon mixing *(releases energy)*

Solution Formation and IMFs
**POLAR / IONIC** dissolve in POLAR:
• ΔH required to overcome IMFs in both the polar/ionic solute and the polar water molecules is quite **LARGE**
• ΔH released due to the interactions between the polar/ionic solute and the polar water molecules is very **LARGE**
• Solute can dissolve because it gets as much energy “back” from the interactions as was required to overcome the IMFs
• Same goes for nonpolar/nonpolar solutions
**Solution Formation and IMFs**

POLAR / IONIC do NOT dissolve in NONPOLAR:

- ΔH released due to interactions between the polar/ionic solute and the nonpolar solvent is quite SMALL
- Solute CANNOT dissolve because the energy required to overcome the IMFs is not provided by the solute-solvent interactions

**Intermolecular Forces**

- Gases have WEAK IMFs and lots of motion
- Solids have STRONG IMFs and no motion
- Liquids have both STRONG IMFs and motion

**IMFs and Properties**

- When considering strength of properties, remember:

  - **Dispersion < Dipole-Dipole < H-Bonding**
    - Linear / Flat molecules = MORE IMFs
    - Spherical / Branched molecules = LESS IMFs
    - If molecules have REALLY large e- cloud (or size), dispersion forces could be strongest!

**Types of Solids**

- **CRYSTALLINE SOLIDS**
  - Highly regular arrangement of their components (IONIC)
  - crystalline

- **AMORPHOUS SOLIDS**
  - Considerable disorder in their structures (GLASS)
  - amorphous
Main Types of Solids

- **MOLECULAR**: orderly 3D structure with discretely bonded molecules at each of its lattice points / Low MP (Ex: water, dry ice, P₄, S₈)

Main Types of Solids

- **IONIC**: orderly pattern of anions and cations; ions occupy lattice points / Very high MP (Ex: NaCl, CaCl₂)
  - Closest Packing → Large particles (anions) get arranged in a way that they are all packed tightly together; Small particles (cations) fit into the holes... Maximizes +/- attraction and minimizes +/+ or -/- repulsions

Main Types of Solids

- **COVALENT NETWORK**: large networks of strong covalent (C or S) bonds (Ex: Quartz, Graphite, Diamond)
  - Diamond → All C are sp³ hybridized making it strong and hard (must break ALL the bonds)
  - Graphite → Forms layers of C atoms arranged in fused sp² hybridized rings (delocalized e- allow it to conduct)

Main Types of Solids

- **ATOMIC**: includes atoms held together by LDF (Ex: Noble gases) and atoms with free-moving e- (Ex: Metals)

Main Types of Solids

- **COVALENT NETWORK**: large networks of strong covalent (C or S) bonds (Ex: Quartz, Graphite, Diamond)
  - Quartz / Silica (SiO₂) → Structure is actually based on a interconnected SiO₄ tetrahedral
  - Glass → Formed by quickly cooling melted silica, giving an amorphous solid / Other compounds added before cooling to give different types

Bonding Models of Metals

- **Electron Sea Model**: regular array of metals in a "sea of electrons" that conduct heat and electricity
- **Bond (Molecular Orbital) Model**: e- travel within molecular orbitals formed by the valence orbitals of the metal cations
**Metal Alloys**
*Mixtures of two or more elements with at least one being a metal*

**SUBSTITUTIONAL**
Host metal atoms are replaced in the lattice by other atoms of similar size
(Ex: Brass, sterling silver, etc.)

**INTERSTITIAL**
Holes in the closest packed metal structure filled by small atoms
(Ex: Steel)

**Phase Changes**
*Change in state*

To change phase, enough energy is needed to allow the molecules to overcome the IMFs!

MORE IMFs = MORE energy needed!

**Vaporization (Evaporation)**
*Molecules of a liquid escaping from the surface to form a gas*

- Always **ENDOTHERMIC**... energy absorbed in order to overcome the IMFs in the liquid
- **Enthalpy of Vaporization (ΔH<sub>vap</sub>):** energy required to vaporize ONE mole of a liquid at 1 atm (heat of vaporization)
- H₂O has a **HIGH** heat of vaporization (40.7 kJ/mol)... lots of energy needed to evaporate!

**Vapor Pressure**
*Pressure of vapor present at equilibrium (rate of condensation equals rate of evaporation)*

- **NONVOLATILE**: liquids with low VPs and high IMFs
- **VOLATILE**: liquids with high VPs and low IMFs
- VP **INCREASES** as temperature **INCREASES**... more energy to escape
- VP **DECREASES** as MM **INCREASES**... more e- mean more polarizable, causing more IMFs! H-bonding is an exception... gives more IMFs so lower VP

**Boiling Point**
*Temperature at which the VP of a liquid equals the atmospheric pressure*

- **Normal boiling point**: temperature at which VP of the liquid equals 1 atm
**Sublimation**

*Substance goes directly from solid to gas*

- Solids have VP, but it's usually very low
- Solids with few IMFs can have high VP and sublime at room temp

**Melting Point**

*Molecules break loose from the lattice points in a solid, changing to a liquid*

- **Enthalpy of Fusion** ($\Delta H_{\text{fus}}$): energy required to convert one mole of a solid to one mole of a liquid
- **Normal melting point**: temperature at which solid and liquid states have same VP where total pressure is 1 atm

**Heating Curve**

$\Delta H_{\text{fus}}$ always less than $\Delta H_{\text{vap}}$ for substance because vaporizing has to overcome ALL IMFs not just some!

- Temperature remains **CONSTANT** during a phase change (PE changes, adding energy = INCR., removing energy = DECR., while KE remains same)
- Chemical bonds are **NOT** being broken during phase changes

**Exceptions**

- Changes of state do not always form at the exact MP and BPs

**SUPERCOOLING**

Rapid cooling allows liquid to remain a liquid at temperature below its FP. Quick temp change does not allow molecules to organize properly (solid will form readily once it happens)

**Exceptions**

- Changes of state do not always form at the exact MP and BPs

**SUPERHEATING**

Rapid heating allows liquid to remain a liquid at temperature above its BP. Quick temp change does not allow enough high energy molecules to accumulate in one place (bubbles very large when form... add boiling chips!)

**Phase Diagrams**

*Represent phases of matter as a function of temperature (x) and pressure (y)... closed systems NOT outdoors!*

- Lines represent when two phases are in equilibrium
- **TRIPLE POINT**: all three phases exist in equilibrium
- **CRITICAL TEMP ($T_c$)**: temp above which the vapor cannot be liquified
- **CRITICAL PRESSURE ($P_c$)**: pressure required to liquefy AT $T_c$
- **CRITICAL POINT**: point defined by $T_c$ and $P_c$
**Phase Diagrams**

**SOLIDS**
Low temp  
High pres

**LIQUIDS**
In between

**GASES**
High temp  
Low pres

*Phase boundaries (lines) represent **EQUILIBRIUM** set of conditions for temp and pressure!!

**Phase Diagrams**

**WATER**
Notice how the solid-liquid line tilts to the left (- slope) since solid is **LESS DENSE** than liquid form

**CARBON DIOXIDE**
Notice how the solid-liquid line tilts to the right (+ slope) since solid is **MORE DENSE** than liquid form