LIQUIDS / SOLIDS / IMFs

Intermolecular Forces (IMFs)
Attractions BETWEEN molecules

- 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

Example: $H_2O$ has stronger IMFs than $CO_2$ due to hydrogen bonding.

London Dispersion Forces (LDFs)

- 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

Example:
- $F_2$ and $Cl_2$ gases... $Br_2$ and $I_2$ are not because they can't overcome the IMFs!
*For the AP Exam... it's the ELECTRONS not the mass that is important!*

London Dispersion Forces (LDFs)

- POLARIZABILITY: ease of distortion of the e- cloud
- MORE ELECTRONS (aka greater mass) = GREATER POLARIZABILITY = STRONGER LDFs for a substance!!!

Example: $F_2$ and $Cl_2$ gases... $Br_2$ and $I_2$ are not because they can't overcome the IMFs!
**London Dispersion Forces (LDFs)**

Directly related to the **polarizability** of each substance... more electrons as you go down the periodic table, so **stronger LDFs lead to solids and higher boiling and melting points**!

**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**

*Notice that H₂O, HF, and NH₃ 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!
Intermolecular Forces

• EXAMPLE:
  Identify the main / strongest type of IMF in each.

*Important to understand molecules that have H-bonding have ALL three IMFs present, while molecules with dipole-dipole have 2 types of IMFs (LDF and D-D) present!!

Liquid State Properties

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

  HIGHER IMFs = GREATER!

Liquid State Properties

• CAPILLARY ACTION: spontaneous rising of a liquid in a narrow tube

  HIGHER IMFs = GREATER!

  -COHESIVE FORCES: between molecules
  -ADHESIVE FORCES: between molecules and container
  -Water... Adhesive GREATER than cohesive (concave meniscus)... Hg would be convex!

Capillary Action

Narrower the tube...
More surface area of glass = HIGHER!

Liquid State Properties

• VISCOSITY: resistance to flow / increases with greater size and lower temps too

  HIGHER IMFs = GREATER!

Liquid State Properties

• VOLATILITY: ability to evaporate readily

  HIGHER IMFs = LESS VOLATILE!
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)

DO NOT SAY “LIKE DISSOLVES LIKE” ON THE AP EXAM!!

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” (and more) 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!

IMFs and Properties

• EXAMPLE: Arrange the following in order of increasing boiling point.
  Ethanol, Propane, Ether
**Intermolecular Forces**

**Types of Solids**

**CRYSTALLINE SOLIDS**

- Highly regular arrangement of their components (IONIC)
- Crystalline
- Amorphous

**AMORPHOUS SOLIDS**

- Considerable disorder in their structures (GLASS)

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**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)

- **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

- **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)

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**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
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 Diagram](image)

Bonding Models of Metals

- **Electron Sea Model**: regular array of metals in a "sea of electrons" that conduct heat and electricity
- **Band (Molecular Orbital) Model**: e- travel within molecular orbitals formed by the valence orbitals of the metal cations

![Bonding Models of Metals Diagram](image)

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 (radii are close)

- Still remains somewhat malleable and ductile with a similar density to the original metal
- Ex: Brass, sterling silver, etc.

![Metal Alloys: Substitutional](image)

**INTERSTITIAL**
Holes in the closest packed metal structure filled by small atoms, typically nonmetals

- Becomes more rigid, less malleable and ductile with a higher density
- Ex: Steel

![Metal Alloys: Interstitial](image)

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!

![Phase Changes Diagram](image)

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** ($\Delta H_{\text{vap}}$): energy required to vaporize ONE mole of a liquid at 1 atm (heat of vaporization)
- $\text{H}_2\text{O}$ 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** ($\Delta H_{fus}$): 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**

- 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).

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 $P_c$

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

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!