



AP Chemistry Ultimate Review Packet

Unit 3: Intermolecular Forces & Properties

Intermolecular Forces

There are several types of **intermolecular forces**:

1. **Ion-ion interactions** – interactions between formally charged ions, strongest intermolecular force, holds together ionic solids
2. **Ion-dipole interactions** – interactions between ions and dipoles, second strongest intermolecular force
3. **Dipole-dipole interactions** – interactions between dipoles, third strongest intermolecular force, electrostatic interactions between neutral polar molecules
4. **Hydrogen bonds** – interactions between NH/OH/FH bond dipoles, technically a very strong dipole-dipole interaction
5. **Dispersion interactions** – between momentary and induced dipoles, weakest intermolecular force
 - any chemical substance is capable of this interaction
 - electron cloud can be slightly skewed to produce momentary dipole
 - when this approaches another particle it can generate an induced dipole
 - these tiny fleeting dipoles can interact briefly and weakly

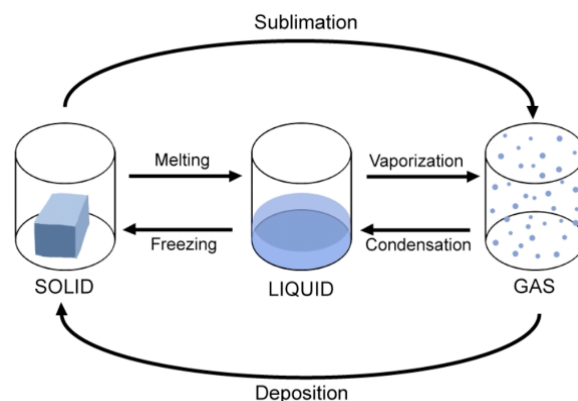
The *strength* of an intermolecular interaction depends on the degree of charge:

(2 formal > 1 formal + 1 partial > 2 partial > momentary/induced)

The Phases of Matter

There are three **phases of matter** relevant to chemistry. Any chemical substance can be in any of these phases depending on temperature/pressure

1. **solid** – particles are very close and not moving with respect to each other
2. **liquid** – particles are very close but moving fluidly past each other
3. **gas** – particles are far apart and in constant motion



Many possible **phase changes**:

- 1) solid → liquid = melting, liquid → solid = freezing
- 2) liquid → gas = vaporization, gas → liquid = condensation
- 3) solid → gas = sublimation, gas → solid = deposition

Phase a substance takes at room temperature and temperatures of phase changes depend on the strength of the intermolecular forces within the substance:

- weaker forces = more easily melt and boil (gas at low temp)
- stronger forces = more difficult to melt and boil (solid or liquid at high temp)

There are a number of **types of solids**:

- 1) crystalline solids – reliable, repeating structure
 - a) molecular solids – solid phase for small molecules (e.g. H₂O)
 - b) network solids – covalent bonds throughout substance (e.g. diamond)
 - c) ionic solids – lattice of cations and anions (e.g. NaCl)
 - d) metallic solids – lattice of metal atoms (e.g. Na, Fe, Ag, etc.)
- 2) amorphous solids – no reliable or repeating structure
 - often the case with large molecules like polymers and gels



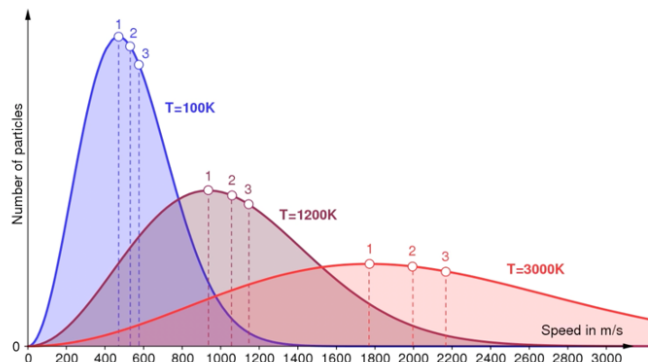
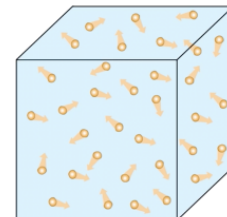
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Ideal Gas Laws

Ideal gases can be described by the postulates of **kinetic molecular theory**:

1. a gas consists of tiny particles always in motion (elastic collisions)
2. gas particles are treated as dimensionless points (negligible volume)
3. pressure is exerted when the particles collide with the container boundary
4. it is assumed that the gas particles do not interact with one another
5. average kinetic energy of the particles is proportional to the temperature in K (this is visible on a Maxwell-Boltzmann distribution)



Calculations can be done with ideal gases using **ideal gas laws**:

- 1) **Boyle's Law** – $P_1V_1 = P_2V_2$
 - pressure and volume are inversely proportional
- 2) **Charles's Law** – $V_1/T_1 = V_2/T_2$
 - volume and temperature are directly proportional
- 3) **Gay-Lussac's Law** – $P_1/T_1 = P_2/T_2$
 - pressure and temperature are directly proportional
- 4) **Combined Gas Law** – $P_1V_1/T_1 = P_2V_2/T_2$
- 5) **Avogadro's Law** – $V_1/n_1 = V_2/n_2$
 - volume and moles are directly proportional
 - leads to molar gas volume, any ideal gas = 22.4 L/mol at STP
- 6) **Dalton's Law of Partial Pressures** (for mixtures of gases)
 - partial pressure of a gas is proportional to its mole fraction
- 7) **Ideal Gas Law** – $PV = nRT$
 - can be used to relate all variables of a gas in isolated state

Ideal gas laws apply when temperature is high and pressure is low

When conditions are not ideal we are dealing with **real gases**:

- described by the van der Waals equation: $(P + an^2/v^2)(V - nb) = nRT$
- correction factors account for volume of particles and interactions between them
 - *high pressures* mean gas is dense and particle volume is not negligible
 - *low temperatures* mean particles move slowly and are able to interact

Solution Chemistry

A mixture in the liquid phase is called a **solution**

- the **solute** is the thing that is getting dissolved (present in lesser amount)
- the **solvent** is the thing doing the dissolving (present in greater amount)
- ionic or covalent solids can dissolve in water, entering the aqueous phase
- if ionic the ions are also dissociating to form ion-dipole interactions

Solution concentration is expressed in **molarity** (M) or **molality** (m)

- M = moles solute per liter of solution (mol/L)
- m = moles solute per kilogram of solvent (mol/kg)



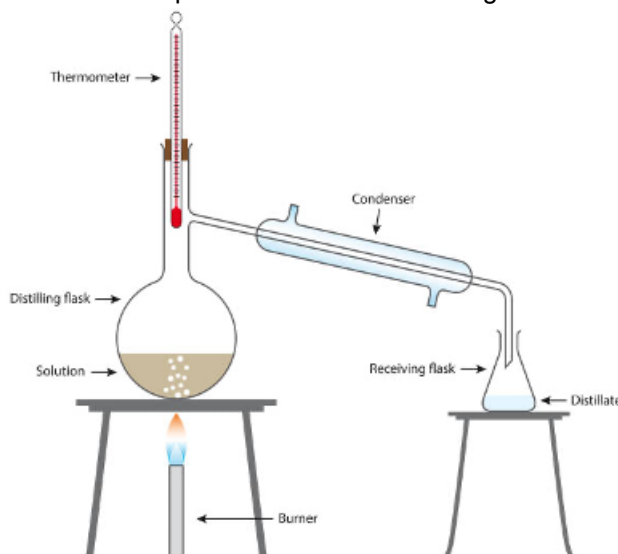
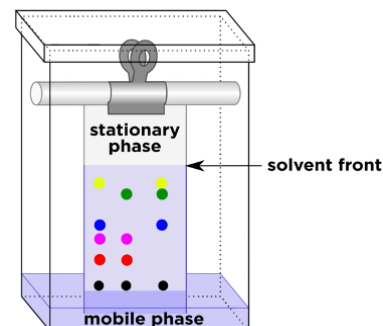
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Separation Techniques

Compounds in a liquid mixture can be separated in several different ways:

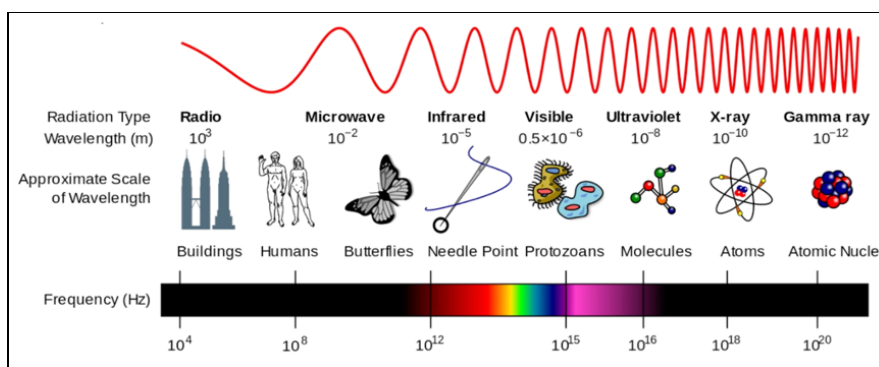
- **Chromatography** separates by virtue of differing polarity
 - solvent travels up plate/paper by capillary action
 - components travel up with the solvent at differing rates
 - more polar = interact with stationary phase better = slower/lower
 - less polar = get pushed along with solvent front = faster/higher
- **Distillation** separates by virtue of differing boiling points
 - mixture is brought to temperature *above* BP of compound 1 but *below* BP of compound 2, so only compound 1 vaporizes
 - vapor enters condenser and liquid is collected in receiving flask



The Electromagnetic Spectrum

Electromagnetic radiation can possess a wide range of frequencies. These are all displayed on an **electromagnetic spectrum**

- microwave radiation is associated with transitions in *molecular rotational levels*
- infrared radiation is associated with transitions in *molecular vibrational levels*
- UV/visible radiation is associated with transitions in *electronic energy levels*



Measuring the way a solution absorbs light of particular wavelengths is useful. This is done with the **Beer-Lambert law** ($A = \epsilon bc$)

A = absorbance (between 0 and 1)

ϵ = molar absorptivity (M/cm)

b = path length (distance light travels through sample in cm)

c = concentration (M)