

#61 Notes Unit 8: Solids/Liquids
Ch. Solids/Liquids

**

Type of Solid (compound)	Type of Bonding (Interparticle Force)	Properties	Examples
Ionic	Ionic	-hard -high melting pts.	NaCl, CaF ₂
Molecular	<i>Covalent:</i>	-soft -low melting points	
	<i>Polar:</i> H-Bonding	→→→→→→→→→→→→→→→→	H ₂ O
	Dipole-Dipole	→→→→→→→→→→→→→→→→	PCl ₃
	<i>Nonpolar:</i> London Dispersion Forces (LDF)	→→→→→→→→→→→→	CO ₂

**Leave room to continue table in a few days!!

I. Ionic Solids

-are made of ions {(+) and (-)} and held together by electrostatic attraction.
 -stable, high melting, but dissolve in polar substances, like H₂O.

3 Types of Closest Packed Structures



Trigonal Holes (space between 3 atoms)
 The smaller (+) ions would fit in between the larger (-) ions in the trigonal hole. Na⁺ (small) and Cl⁻ (large)



Tetrahedral Holes (space between 4 atoms)



Octahedral Holes (space between 6 atoms)

trigonal holes << tetrahedral holes << octahedral holes
 (smallest) (largest)

II. Molecular Solids

-are covalently bonded. (nonmetals with nonmetals or **"H"**)

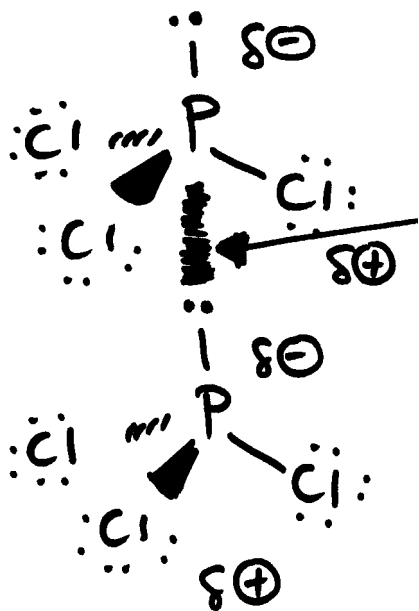
Van der Waals Forces:

-are forces **between** molecules (**Intermolecular** forces).

{ **Intramolecular** forces are bonds **within** a single molecule. }

A) Dipole-Dipole Forces:

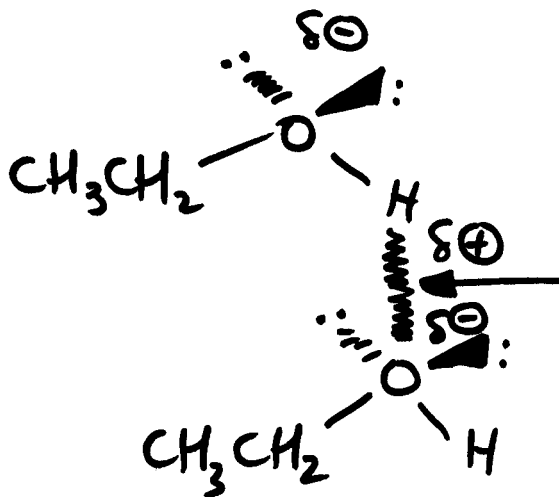
-are between 2 polar molecules



Dipole-dipole force = the $\delta(+)$ end of one molecule is attracted to the $\delta(-)$ end of another molecule.

Hydrogen Bonding (** A special type of dipole/dipole force):

This occurs in molecules when a "H" is bonded to a highly electronegative atom (N, O, F). The hydrogen bond is between the "H" in one molecule and the N, O, or F in the other molecule.



Hydrogen bond

*These are stronger than dipole/dipole forces, since the N, O, F are so electronegative that the partial charges are nearly whole.

B) London Dispersion Forces (LDF)

-are between 2 nonpolar molecules.

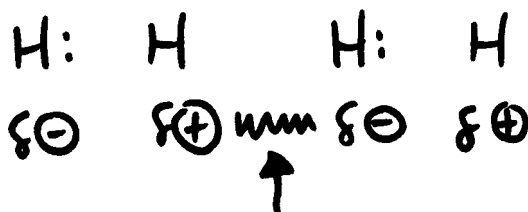
H—H For a split second the electrons happen to be closer to one of the H's.
In this split second:



Now the electrons in another nearby H₂ molecule will either be attracted or repelled by the temporary dipole.

Temporary dipole

As these electrons are attracted or repelled, an induced dipole will form.



Temporary Dipole | LDF Force | Induced Dipole

The attraction between the 2 molecules is the LDF Force.

**The larger the molecule, the stronger the LDF Force, since there are more electrons to get out of balance (creating stronger dipoles).

**The longer the molecule, the stronger the LDF Force, since longer molecules will tangle and stick together.

Interparticle Forces:

Strongest

Ionic

Na⁺Cl⁻

(+) and (-) ions

>

H-Bonding

-OH

HF

-NH₂, -NH, NH₃

H with N,O,F

>

Dipole-Dipole

polar

↑

>

Weakest

LDF

nonpolar

↑

must be drawn out like last chapter

Remember polarity lab: CH₃CH₂CH₂CH₂CH₃ pentane (nonpolar, so LDF)

Ex. 1) Compound

a) CH₃CH₂CH₂NH₂

b) Mg(NO₃)₂

Type of Interparticle Force

H-Bond

Ionic

Type of Solid

Molecular

Ionic

c) SF₄

(seesaw, so polar) **Dipole-Dipole**

Molecular

d) CH₃CH₃

(nonpolar) **LDF**

Molecular

#62 Notes III. Liquid Forces

-are intermolecular forces too weak to hold the atoms together in a solid form.

1) Adhesive Forces

-are between a liquid and its container.

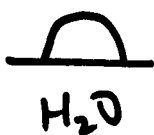


a) Capillary action: is the rising of a liquid in a small tube.
(attraction to glass)

2) Cohesive Forces

-are intermolecular forces between liquid molecules.

a) Surface Tension: is the resistance of a liquid to increase its surface area.



H₂O



alcohol

(Hg would be an almost round drop.)

H₂O has higher surface tension (it sticks together better), than alcohol

b) Viscosity: is a measure of a liquid's resistance to flow

Oil has higher viscosity (it sticks together better) than H₂O.

Cohesive Forces:

**The greater the interparticle force (the more sticky it is), the:

-greater the surface tension (liquid sticks).

-greater the viscosity (liquid sticks).

-greater the boiling point

(liquid sticks, hard to escape as a gas).

-greater the melting point (freezing point)

(solid sticks, hard to melt).

-weaker the vapor pressure

(liquid sticks, less liquid evaporating to form a gas).

Greatest Interparticle Force (stickiest)



Ionic
(+)/(-) ions

> metallic
metals
(Fe, Pu)

> H-Bonding
H with N,O,F

> Dipole-Dipole
polar

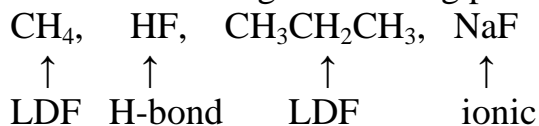
Weakest



LDF*
nonpolar

*(for LDF: long chains (tangle) and atoms with more electrons, increase the LDF interparticle force)

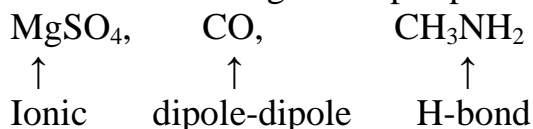
Ex.1) Which has the highest melting point and the lowest melting point?



highest m.p. = hard to melt (stickiest, strongest force) = NaF (ionic),

lowest m.p. = easy to melt (least stickiest, weakest force) = CH₄ (LDF, with shortest chain, less tangles)

Ex. 2) Which has the highest vapor pressure?



highest v.p. = a lot evaporates (least stickiest, weakest force) = CO (dipole-dipole)

IV. Solids

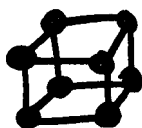
Crystals: are rigid bodies in which the particles are arranged in a repeating pattern.

Amorphous Solids: lack this regular repeating pattern. (plastics, butter)

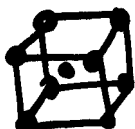
The crystal's pattern depends on the type of bonding and the size of the particles.

The smallest repeating unit is the unit cell.

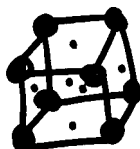
3 simplest unit cells:



Simple cubic



Body centered cubic



Face centered cubic

The structure of solids can be determined by x-ray diffraction.

A crystal is rotated in a x-ray beam and the constructive/destructive interference patterns are interpreted by a computer, using the Bragg equation.

$$n \lambda = 2 d \sin \theta \quad n = \text{order (1, 2, \dots etc.)}$$

λ = wavelength in meters
 d = interplanar distance between atoms
 θ = angle of incidence or reflection

$1 \text{ \AA} = 1 \times 10^{-10} \text{ m}$
 (angstrom)

*Continuation of table from first day of notes.

Type of Solid (<u>compound</u>)	Type of Bonding (<u>Interparticle Force</u>)	Properties _____	Examples _____
Metallic	non-directional covalent (bend)	-wide range of hardness and melting points -conductors	Ag, Fe Brass: (Zn & Cu)
Network	directional covalent (don't bend)	-hard -high melting pts.	atomic: C, metalloids (Si, B, Ge) molecular: SiO ₂ (sand), S ₈ , P ₄ C – diamond, graphite, charcoal
Group VIII	LDF	-low melting pts.	noble gases

Ex. 3) Which has the lowest boiling point?

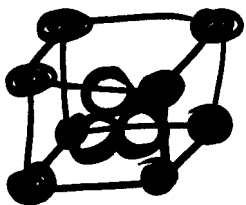
Pd,	CH ₃ CH ₂ OH,	SrSO ₃ ,	PF ₃
↑	↑	↑	↑
metallic	H-bond	ionic	polar (trigonal pyramidal) dipole-dipole

lowest b.p. = easy to boil (least stickiest, weakest force) = PF₃ dipole-dipole

#63 Notes V. Metallic

1) 2 Types of closest packing of metals

-Hexagonal closest packing (ABAB)

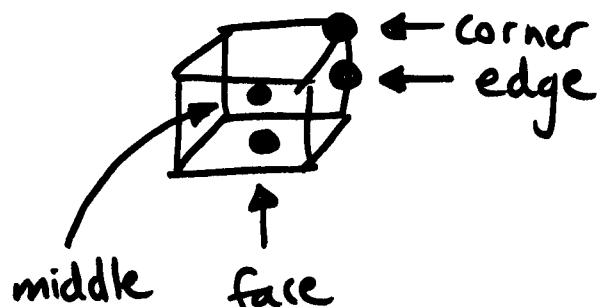


Hexagonal prism: body centered, but with parts of 3 atoms in the center.

-Cubic closest packing (ABCABC)

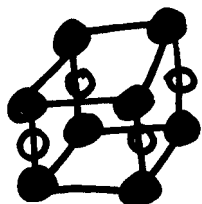
Face centered (alkali metals)

Positions of atoms when packed (for counting number of atoms in a unit cell):



Corner = 1/8 of an atom
Edge = 1/4 of an atom
Middle = 1 atom
Face = 1/2 of an atom

Ex. 1)

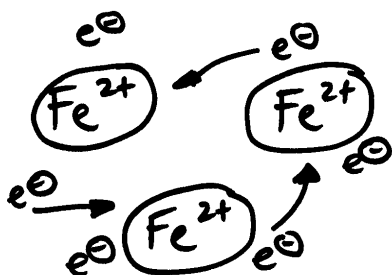


Na: 8 corners (1/8) = 1 atom Na
Cl: 4 edges (1/4) = 1 atom Cl,
so **NaCl**
if Na_3Cl_3 reduce to Na_1Cl_1

2) Bonding Theories:

Nondirectional bonds: since they tend to stay bonded as long as they are in contact.
(malleability)

A) Electron Sea Model



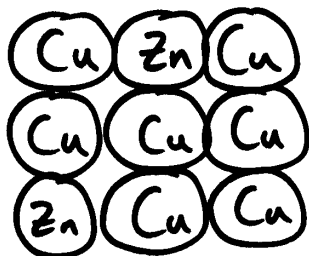
-nondirectional bonding
-sea of electrons free to move around in currents
(conduction of electricity)

B) Band or Molecular Orbit Model

-same, except the e^- are transferred between molecular orbitals formed from the valence atomic orbitals (d-orbitals overlap in conduction bands).

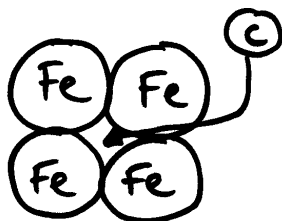
3) Alloys

-mixtures of metals



Brass

a) substitutional alloy: host metal is replaced by a metal of a similar size.



Steel

b) interstitial alloy (between atoms): holes of the closest packed structure are filled by smaller atoms

VI. Network Solids

-atomic: C, metalloids (Si, B, Ge, ...)

-molecular: S_8 , P_4 , SiO_2

-strong directional covalent bonds to make giant molecules

C -diamond, graphite, charcoal

Si_xO_y -silicates, includes glass, quartz

Glass = SiO_2 melted at $1600^\circ C$ and cooled rapidly.

$Al_xSi_yO_z$ -aluminosilicates (ceramics) & Al_2O_3 compounds

A) Semiconductors (semimetals, metalloids)

-somewhat conduct electricity, since there is a small gap between filled and unfilled orbitals that the e^- can jump.

1) Doped Semiconductors: a mixture of elements.

a) n-type: elements in the material are replaced by atoms that have extra valence e^- to increase conduction. (Te added to Si)

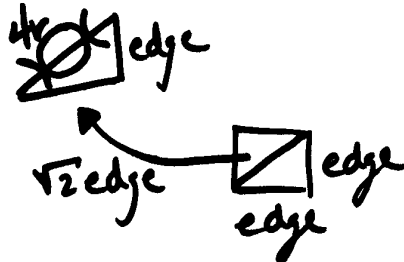
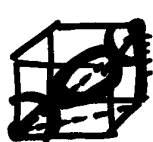
b) p-type: doped (replaced) with elements that have fewer valence e^- to decrease conduction. (Si added to Te)

#64 Notes VII. Group VIII

- noble gases
- low melting points & boiling points
- nonpolar (atomic), so LDF forces

VIII. Body Centered/ Face Centered Problems

Ex. 1) Molybdenum is body centered cubic and has an atomic radius of 139 pm. Calculate its density.



$$\begin{array}{l} 8 \text{ corners } (1/8) = 1 \\ \hline 1 \text{ middle} = 1 \\ \hline 2 \text{ atoms} \end{array}$$

1) Find edge:

$$\begin{aligned} (4r)^2 &= (\text{edge})^2 + (\sqrt{2} \text{ edge})^2 \\ [(4)(r)]^2 &= 3 (\text{edge})^2 \\ 16 \cdot r^2 &= 3(\text{edge})^2 \\ (16) (139 \text{ pm})^2 &= 3(\text{edge})^2 \\ 309136 &= 3(\text{edge})^2 \\ 321 \text{ pm} &= \text{edge} \end{aligned}$$

2) Find volume in cm^3/mol :



$$V = \ell \cdot w \cdot h = (321 \text{ pm})^3 = 3.3078 \times 10^7 \text{ pm}^3$$

$$\frac{3.3078 \times 10^7 \text{ pm}^3}{2 \text{ atoms}} \left| \frac{(1 \times 10^{-12} \text{ m})^3}{(1 \text{ pm})^3} \right| \left| \frac{(1 \text{ cm})^3}{(1 \times 10^{-2} \text{ m})^3} \right| \left| \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} \right| = 9.96 \text{ cm}^3/\text{mol}$$

3) Find density:

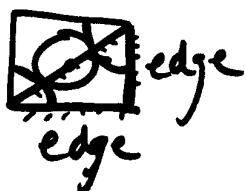
$$D = m/v = \frac{95.94 \text{ g/mol (molybdenum from periodic table)}}{9.96 \text{ cm}^3/\text{mol}} = \mathbf{9.63 \text{ g/cm}^3}$$

Differences for Face Centered (= Cubic Closest Packing):

$$8 \text{ corners } (1/8) = 1$$

$$6 \text{ faces } (1/2) = \underline{3}$$

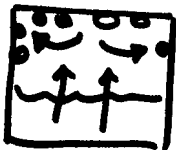
$$4 \text{ atoms}$$



$$\begin{aligned} (4r)^2 &= (\text{edge})^2 + (\text{edge})^2 \\ 16r^2 &= 2 (\text{edge})^2 \end{aligned}$$

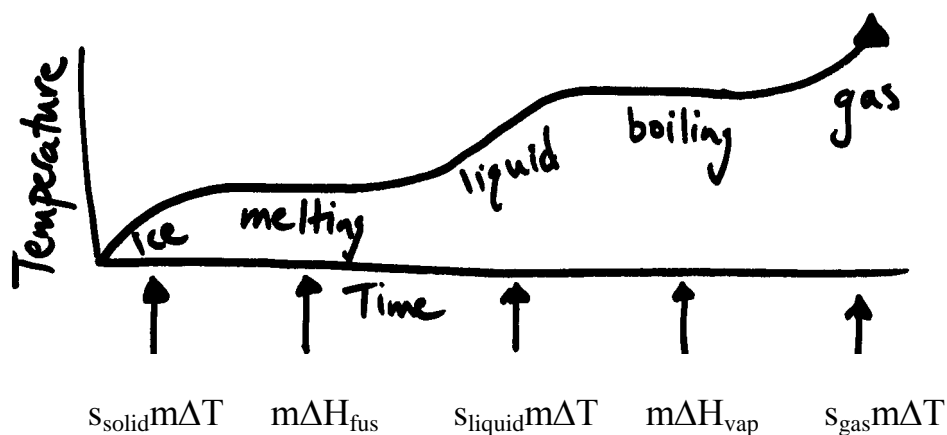
IX. Vapor Pressure

-is caused by gas escaping the surface of a solid or a liquid.



Equilibrium: when the number of particles turning from gas to liquid (condensation) equals the number of particles turning from liquid to gas (evaporation).

* *evaporation is affected by temperature, increase temperature, increase vapor pressure.*



ΔH_{fusion} : Enthalpy of Fusion is the energy required to melt 1 mol of a solid at 1 atm of pressure.

* *If the solid and liquid are at the same temperature at the melting point, the Kinetic Energy would be the same, the velocity of the particles would be the same, the evaporation rates would be the same, the vapor pressures would be the same!*



Melting Point is the temperature at which the solid and the liquid have the same vapor pressure. (same Kinetic Energy)

$\Delta H_{\text{vaporization}}$: Enthalpy of Vaporization is the energy required to vaporize 1 mol of a liquid at 1 atm of pressure.

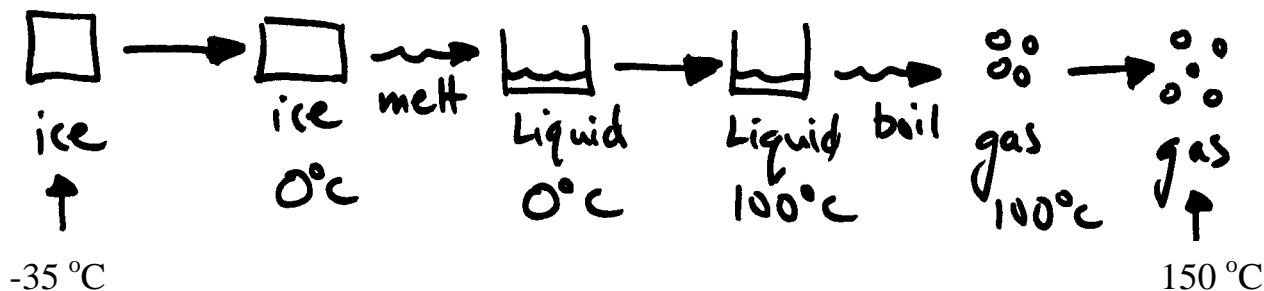
****Boiling Point** is the temperature at which the liquid's vapor pressure equals the atmospheric pressure. {normal b.p. at standard pressure (1 atm, 760 torr)}

For H₂O: $s_{\text{solid}} = 2.1 \text{ J}/\{\text{g}\cdot^{\circ}\text{C}\}$ $\Delta H_{\text{fus}} = 6.02 \text{ kJ/mol}$

$s_{\text{liquid}} = 4.2 \text{ J}/\{\text{g}\cdot^{\circ}\text{C}\}$ $\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}$
(4.184 from energy chapter)

$s_{\text{gas}} = 2.0 \text{ J}/\{\text{g}\cdot^{\circ}\text{C}\}$

Ex. 1) How much energy does it take to convert 1.75 kg ice at -35°C to steam at 150°C ?



$$\frac{1.75 \text{ kg}}{1 \text{ kg}} \left| \frac{1 \times 10^3 \text{ g}}{1 \text{ kg}} \right. = 1.75 \times 10^3 \text{ g H}_2\text{O} \left| \frac{1 \text{ mol}}{18.015 \text{ g}} \right. = 97.1 \text{ mol H}_2\text{O}$$

$$q = s_{\text{solid}}m\Delta T + m\Delta H_{\text{fus}} + s_{\text{liquid}}m\Delta T + m\Delta H_{\text{vap}} + s_{\text{gas}}m\Delta T$$

$$q = (2.1 \text{ J}/\text{g}\cdot^{\circ}\text{C}) (1.75 \times 10^3 \text{ g}) (0 - (-35^{\circ}\text{C})) + (97.1 \text{ mol}) \left(\frac{6.02 \text{ kJ}}{\text{mol}} \left| \frac{1 \times 10^3 \text{ J}}{1 \text{ kJ}} \right. \right)$$

$$+ (4.2 \text{ J}/\text{g}\cdot^{\circ}\text{C}) (1.75 \times 10^3 \text{ g}) (100 - 0^{\circ}\text{C}) + (97.1 \text{ mol}) \left(\frac{40.7 \text{ kJ}}{\text{mol}} \left| \frac{1 \times 10^3 \text{ J}}{1 \text{ kJ}} \right. \right)$$

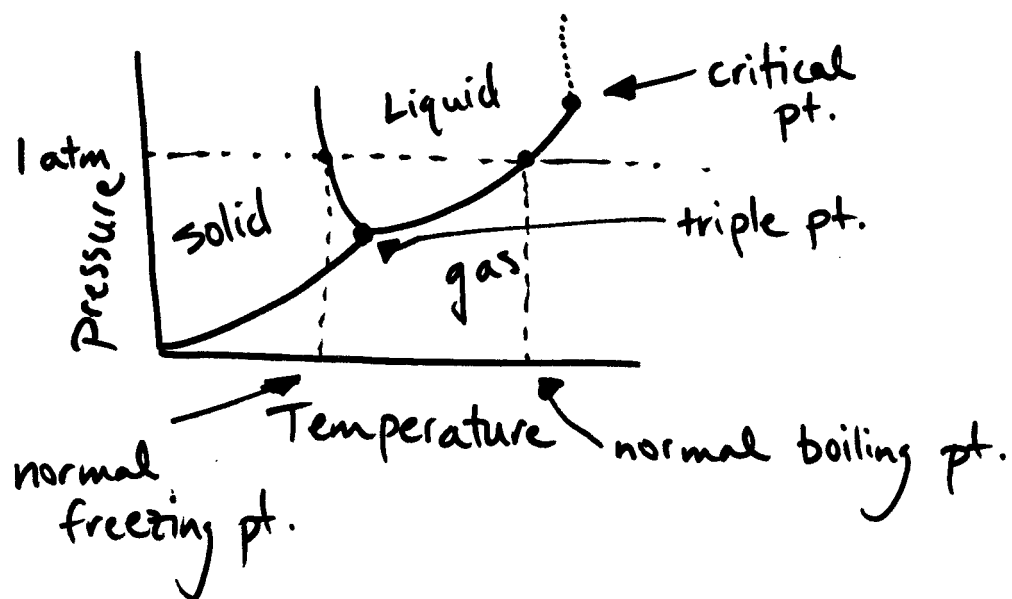
$$+ (2.0 \text{ J}/\text{g}\cdot^{\circ}\text{C}) (1.75 \times 10^3 \text{ g}) (150 - 100^{\circ}\text{C})$$

$$q = 128625 \text{ J} + 584542 \text{ J} + 735000 \text{ J} + 3951970 \text{ J} + 175000 \text{ J}$$

$$q = 5575137 \text{ J} = \mathbf{5.6 \times 10^6 \text{ J}}$$

X. Phase Diagrams

-represent the phases of a substance as a function of T & P.

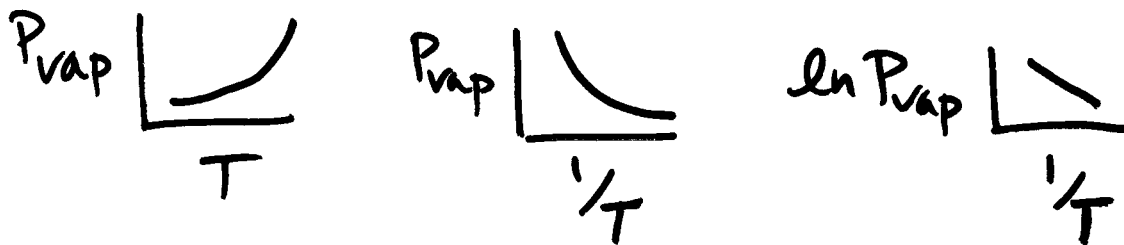


Triple Point: is the point where all 3 states are in equilibrium.

Critical Point: is the highest temperature that a gas can still be compressed into a liquid, (liquefaction).

#66 Notes XI. Vapor Pressure & Temperature

- vapor pressure increases as temperature increases.



If $y = mx + b$, then $\ln P_{vap} = \text{slope} (1/T) + \text{constant}$

↑
Slope = $-\Delta H_{vap} / R$ where $R = 8.31 \text{ J/mol}\cdot\text{K}$

$$\ln P_{vap} = \frac{-\Delta H_{vap}}{R} (1/T) + \text{constant}$$

(You can plug in the values and solve for the constant.)

Leave the units of P in mm or torr, when graphing and calculating the slope.

To graph: Given $\frac{P}{T} \rightarrow \frac{\ln P}{T \text{ in Kelvin}} \mid \frac{1}{T}$

Then graph and calculate the slope: $\frac{(y_2 - y_1)}{(x_2 - x_1)} = \frac{-\Delta H_{vap}}{R}$

And solve for ΔH_{vap}

For 2 Temperatures:

$$\ln P + \frac{\Delta H_{vap}}{RT} = C$$

$$\begin{array}{c} \leftarrow \\ \uparrow \\ \ln P_1 + \frac{\Delta H_{vap}}{RT_1} = \ln P_2 + \frac{\Delta H_{vap}}{RT_2} = C \\ \downarrow \\ \rightarrow \end{array}$$

$$\ln P_1 - \ln P_2 = \frac{\Delta H_{vap}}{RT_2} - \frac{\Delta H_{vap}}{RT_1}$$

$$\ln (P_1 / P_2) = \frac{\Delta H_{vap}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

Clausius-Clapeyron

* $\Delta H_{vap} = \text{kJ} \rightarrow \text{J}$, $R = 8.31 \text{ J/mol}\cdot\text{K}$

Ex. 1) Find the vapor pressure of H₂O at 50 °C, if $\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}$.

-no constant "C", so we must think of another set of values to use.

We know when H₂O normally boils, *show Phase Diagram!*

$T_{\text{bp}} = 100 \text{ °C}$ at $P = 1 \text{ atm} = 760 \text{ torr}$

(by definition: vapor pressure = atmospheric pressure at the boiling pt.)
(standard pressure)

$$\begin{array}{ll} P_1 = 1 \text{ atm} & P_2 = ? \\ T_1 = 100 \text{ °C} = 373 \text{ K} & T_2 = 50 \text{ °C} = 323 \text{ K} \end{array}$$

$$\ln (P_1 / P_2) = \frac{\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

$$\ln (1.00 \text{ atm} / P_2) = \frac{(40.7 \text{ kJ} \mid 1000 \text{ J})}{\text{mol} \mid 1 \text{ kJ}} \left(\frac{1}{323 \text{ K}} - \frac{1}{373 \text{ K}} \right)$$

(8.31 J/mol•K)

$$\ln (1.00 / P_2) = 4898 (0.003095975 - 0.002680965)$$

$$\ln (1.00 / P_2) = 4898 (0.00041501)$$

$$e^x (\ln (1.00 / P_2) = 2.03)$$

$$1.00 / P_2 = 7.63$$

$$1.00 = 7.63 (P_2)$$

$$\mathbf{0.131 \text{ atm} = P_2} \quad (= 99.5 \text{ torr})$$

**** You can always use the normal boiling point (T_{bp} for H₂O = 100 °C) at standard pressure ($P_1 = 1 \text{ atm} = 760 \text{ torr}$)!**

I. Concentration

$$\text{Molarity (M)} = \frac{\text{mols solute}}{\text{L solution}}$$

$$\text{molality (m)} = \frac{\text{mols solute}}{\text{kg solvent}}$$

$$\text{mass \%} = \frac{\text{mass solute}}{\text{mass solution}} \times 100$$

$$\text{mol fraction of solute (X)} = \frac{\text{mols solute}}{\text{mols solute} + \text{mols solvent}}$$

solute: usually the smaller quantity

solvent: usually the larger quantity, but water is always the solvent, if present

Ex. 1) Calculate the molarity, molality and mol fraction of a 20.% solution of ammonia, NH_3 , with a density of 0.88 g/cm^3 .

1) Use mass % to find grams of each substance.

$$\text{mass \%} = \frac{\text{mass solute}}{\text{mass solution}} \times 100$$

$$20\% = \frac{?}{?} \times 100 \rightarrow \frac{20 \text{ g solute (NH}_3\text{)}}{100 \text{ g solution (NH}_3\text{ + H}_2\text{O)}}$$

$$\begin{array}{l} 100 \text{ g solution} \\ - 20 \text{ g NH}_3 \text{ (solute)} \\ \hline 80 \text{ g H}_2\text{O (solvent)} \end{array}$$

2) Get mols of each substance.

$$\frac{20 \text{ g NH}_3}{17.031 \text{ g}} \left| \frac{1 \text{ mol}}{17.031 \text{ g}} \right. = 1.17 \text{ mol NH}_3 \text{ (solute)}$$

$$\frac{80 \text{ g H}_2\text{O}}{18.015 \text{ g}} \left| \frac{1 \text{ mol}}{18.015 \text{ g}} \right. = 4.44 \text{ mol H}_2\text{O (solvent)}$$

3) Find L of solution.

$$D = m/v \quad 0.88 \text{ g/cm}^3 = \frac{100 \text{ g solution}}{V} \quad \leftarrow \text{always } 100 \text{ g of solution}$$

$$V = \frac{114 \text{ cm}^3}{1 \text{ cm}^3} \left| \frac{1 \text{ ml}}{1 \text{ cm}^3} \right| \left| \frac{1 \times 10^{-3} \text{ L}}{1 \text{ ml}} \right| = 0.114 \text{ L solution}$$

4) Find kg of solvent.

$$\frac{80 \text{ g H}_2\text{O}}{1 \times 10^3 \text{ g}} \left| \frac{1 \text{ kg}}{1 \times 10^3 \text{ g}} \right| = 0.080 \text{ kg H}_2\text{O (solvent)}$$

$$M = \frac{\text{mols solute}}{\text{L solution}} = \frac{1.17 \text{ mol NH}_3}{0.114 \text{ L solution}} = 10.3 = \mathbf{10. M}$$

$$m = \frac{\text{mols solute}}{\text{kg solvent}} = \frac{1.17 \text{ mol NH}_3}{0.080 \text{ kg H}_2\text{O}} = 14.6 = \mathbf{15 \text{ mol/kg}}$$

$$X = \frac{\text{mols solute}}{\text{mols solute} + \text{mols solvent}} = \frac{1.17 \text{ mol NH}_3}{1.17 \text{ mol NH}_3 + 4.44 \text{ mol H}_2\text{O}} = \mathbf{0.21}$$

**If given ml and density of each, solve for grams of each, then mols of each etc.

*largest quantity will be solvent

III. Solubility Factors

-polarity, pressure, temperature, *affecting speed only: surface area (crushing solid), stirring*

A) Polarity

****like dissolves like**

Polar (hydrophilic) dissolves in polar (H₂O).

The more concentrated the charge, the better it dissolves (big charge/small radius) and the more polar the molecule, the better it dissolves .

Nonpolar (hydrophobic) dissolves in non polar (CCl₄, hexane, fat soluble proteins)

B) Henry's Law

The amount of gas dissolved in a solution is directly proportional to the pressure of the gas above the solution.

****The higher the pressure on the liquid, the higher the concentration of the gas in the liquid.**

(Ex. soda: CO₂)

$C = k P$ $k = \text{constant, } C = \text{concentration}$

This does not work well for gases that dissociate in water, like $\text{HCl} \rightarrow \text{H}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$

C) Temperature

As water temperature increases, the solubility of a solid increases, but the solubility of a gas decreases. High Temperature, a lot of solid dissolves, little gas dissolves.

Thermal Pollution: Industry releasing hot water, decreases the amount of O₂, fish die.

Ex. 1) Which solvent, H₂O or CCl₄, does each of the following dissolve in best?

- | | | |
|----------------------------------------------------|---------------------------------------|------------------------|
| a) MgCl ₂ | ionic, charged ions dissolve in water | H₂O |
| b) CH ₃ CH ₂ CH ₃ | nonpolar | CCl₄ |
| c) CH ₃ CH ₂ OH | H-bond, polar | H₂O |
| d) PCl ₅ | nonpolar | CCl₄ |

**** hexane is nonpolar, like CCl₄!**

Ex. 2) Which is more soluble in H₂O?

- a) CH₃CH₂CH₃ or HOCH₂CH₂OH

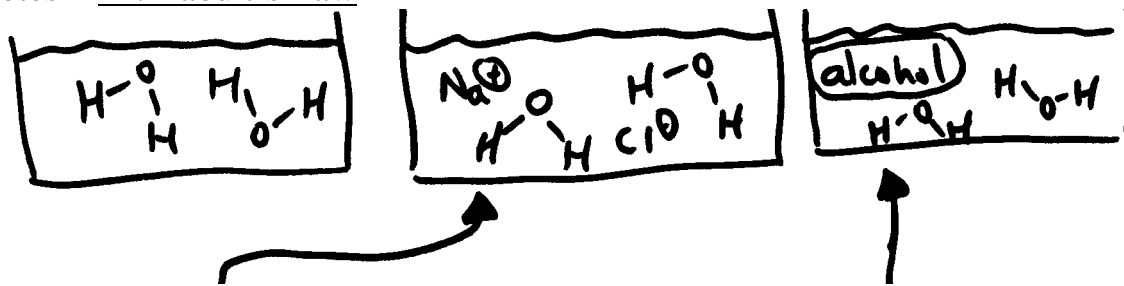
↑

2 H-bonds (more polar)

- b) NaCl or KCl

↑

Na: smaller atom, more concentrated charge



Lowest Vapor Pressure

The nonvolatile salt particles will occupy some of the solvent's surface area, making it harder for the water to escape, decreasing the vapor pressure.

Highest Vapor Pressure

This will have a higher vapor pressure, since the alcohol will evaporate faster than water (alcohol pressure + water pressure).

The amount of solute vs. the amount of solvent and the relative vapor pressures will determine the vapor pressure of the overall solution.

$$P_{\text{solution}} = X_{\text{solvent}} P_{\text{solvent}} + X_{\text{solute}} P_{\text{solute}} \quad \text{*If the solute is nonvolatile, } P_{\text{solute}} = 0.$$

This is for an ideal solution, where all interactions in the solution are equal: solute vs. solute = solute vs. solvent = solvent vs. solvent.

Ex. 1) What is the vapor pressure of a solution containing 219 g pentane (C_5H_{12}) and 50.1 g hexane (C_6H_{14}), if the vapor pressure of pentane is 426 torr at 20°C and hexane is 121 torr at 20°C ?

$$\frac{219 \text{ g } \text{C}_5\text{H}_{12}}{75.15175 \text{ g}} \left| \frac{1 \text{ mol}}{75.15175 \text{ g}} \right. = 3.04 \text{ mol } \text{C}_5\text{H}_{12}$$

$$\frac{50.1 \text{ g } \text{C}_6\text{H}_{14}}{86.1789 \text{ g}} \left| \frac{1 \text{ mol}}{86.1789 \text{ g}} \right. = 0.581 \text{ mol } \text{C}_6\text{H}_{14}$$

$$X_{\text{C}_5\text{H}_{12}} = \frac{3.04 \text{ mol } \text{C}_5\text{H}_{12}}{3.04 \text{ mol} + 0.581 \text{ mol}} = 0.840$$

$$X_{\text{C}_6\text{H}_{14}} = 1 - 0.840 = 0.160$$

(or $X_{\text{C}_6\text{H}_{14}} = \frac{0.581 \text{ mol } \text{C}_6\text{H}_{14}}{3.04 \text{ mol} + 0.581 \text{ mol}} = 0.160$)

$$P_{\text{solution}} = X_{\text{C}_5\text{H}_{12}} P_{\text{C}_5\text{H}_{12}} + X_{\text{C}_6\text{H}_{14}} P_{\text{C}_6\text{H}_{14}}$$

$$P_{\text{solution}} = (0.840)(426 \text{ torr}) + (0.160)(121 \text{ torr})$$

$$P_{\text{solution}} = 358 \text{ torr} + 19.4 \text{ torr} = \mathbf{377 \text{ torr}}$$

If given ml and density, find grams, then mols.

In nonvolatile, $P_{\text{solute}} = 0$, so equation is $P_{\text{solution}} = X_{\text{solvent}} P_{\text{solvent}}$ (usually H_2O)

If given P_{solution} , then solve for P_{solvent} or mol fraction.

$$0.430\text{ }^{\circ}\text{C} = m (1.86\text{ }^{\circ}\text{C}\cdot\text{kg}/\text{mol})$$

$$0.231\text{ mol}/\text{kg} = m$$

$$\text{molality} = \text{mol solute} / \text{kg solvent}$$

$$\text{mol} = \text{mass} / \text{molar mass}$$

$$\text{molality} = \frac{(\text{mass solute})}{(\text{molar mass solute}) (\text{kg solvent})}$$

$$\frac{861\text{ g H}_2\text{O}}{1 \times 10^3\text{ g}} = 0.861\text{ kg}$$

$$(0.231\text{ mol}/\text{kg}) = \frac{(8.20\text{ g})}{(\text{molar mass solute}) (0.861\text{ kg H}_2\text{O})}$$

$$\text{mm} = 41.2\text{ g}/\text{mol}$$

V. Osmotic Pressure

-pressure needed to stop the flow between 2 concentration gradients.

$$PV = nRT$$

$$P = \frac{n}{V} RT$$

$$\pi = i M R T \quad \pi = \text{osmotic pressure (atm)}$$

$$M = \text{molarity (mol/L)}$$

$$R = 0.08206\text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$$

$$T = \text{temperature (K)}$$

i = Van't Hoff Factor = basically the number of ions, but solids don't always break down completely, so " i " will usually be less than the number of ions.

For example:	<u>expected</u>	<u>observed</u> in experiments (found in a table in the book)
MgCl ₂	$i = 3$ ions	$i = 2.7$
FeCl ₃	$i = 4$ ions	$i = 3.4$

Use the expected, unless given the observed (if nonionic, then use 1).

VI. Colloids

-are tiny particles suspended in a medium.

Tyndall Effect: is the scattering of light by these particles.

Brownian Motion: is the random movement of these particles.

Aerosol (liquid or solid in a gas): fog, smoke, dust, aerosol sprays.

Foam (gas in a liquid or solid): Styrofoam, shaving cream, whipped cream.

Emulsion (solid or liquid in a liquid): solids: butter, cheese

liquids: milk, mayonnaise

End of Notes (Assignments #71-72 are Review Assignments. There are no notes for these assignments.)