

Intermolecular Forces and
Solution Chemistry Review

Name Key

- 1) Use the KMT and a discussion on IMFs to explain the molecular activity of matter in its three phases: solid, liquid, and gas.

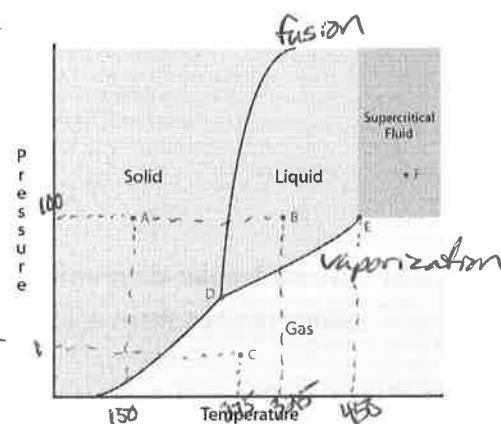
Solid - Rigid structure formed by a high % of IMF attractions intact. Matter is still in motion, per KMT, but movement is highly restricted due to attraction b/w neighboring particles.

Liquid - Fluid nature obtained by decreasing the frequency of IMF attractions. Matter has more freedom of motion, allowing for flowing particle movement.

Gas - All IMF attraction is absent (ideal gas), allowing for particles to expand and fill the volume of their container.

- 2) Explain the shape of the fusion and vaporization curves below using your understandings from #1.

- Both curves bend exponentially at P/T extremes (high P for fusion curve, high T for vaporization curve) showing decreased influence of temp at high pressure & decreased influence of pressure at high temps.



- Be able to predict curve shifts based upon polarity (highly polar = fusion curve shifts right, vaporization down)

3) Assume points A, B, C, D, E, F represent the following temperatures/pressures.

Data Point	Temperature (K)	Pressure (kPa)
A	150	100
B	325	100
C	275	1
D	250	50
E	450	98
F	550	150

Describe what type of matter would be observed under the following conditions:

a. A sample of the substance is heated beyond 450K.

Gas, unless under pressure > 100 kPa (super critical fluid)

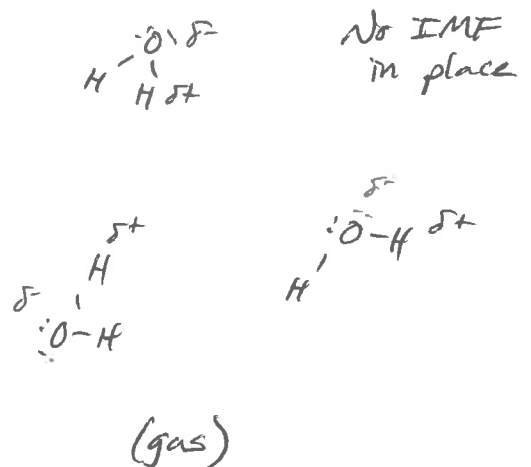
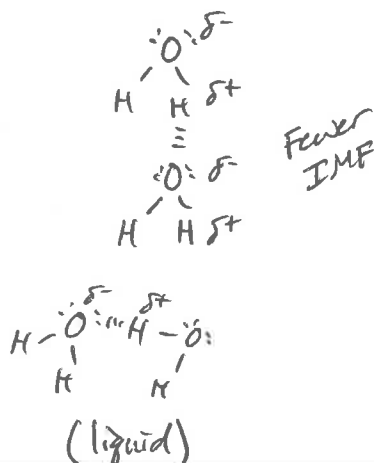
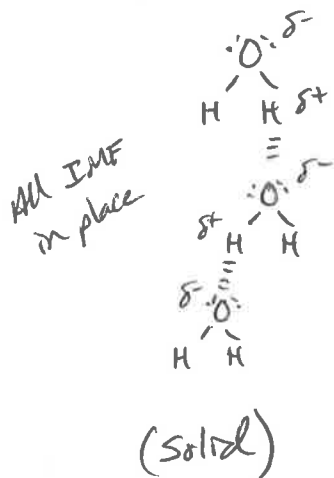
b. A sample is maintained at 340K and 40kPa.

Liquid

c. A sample is maintained at 100kPa and 255K (assume the data point falls on the curve).

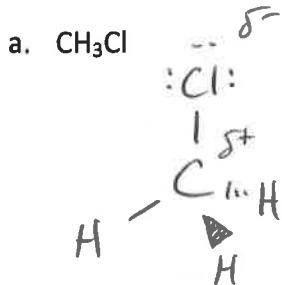
Liquid / solid

4) Use molecular diagrams of H₂O to show what happens at the molecular level during a phase change from solid to liquid and from liquid to gas.



- 5) Identify the type of bond that would occur between the elements in the following molecules and give the ionic characteristic value for the bonds.

$$I = 1 - \exp(-0.25(EN_A - EN_B)^2)$$

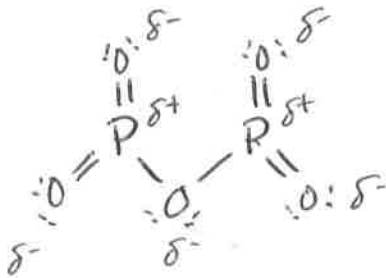


$$I_{\text{C-Cl}} = 1 - e^{\frac{-(3.5-2.5)^2}{4}}$$

$$= 1 - 0.779$$

$$= 0.221 \quad (22.1\% \text{ Ionic Character})$$

- b. P_2O_5

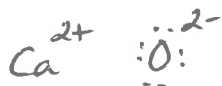


$$I_{\text{P-O}} = 1 - e^{\frac{-(3.5-2.1)^2}{4}}$$

$$= 1 - 0.613$$

$$= 0.387 \quad (38.7\% \text{ Ionic})$$

- c. CaO



$$I_{\text{Ca-O}} = 1 - e^{\frac{-(3.5-1.0)^2}{4}}$$

$$= 1 - 0.210$$

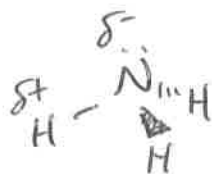
$$= 0.790 \quad (79.0\% \text{ Ionic})$$

- 6) Explain why polarity arises across polar covalent and ionic bonds. Use a Lewis dot diagram of KBr and ammonia to support your answer.

Uneven e^- attraction (ΔEN) causes e^- to be shared unevenly b/w atoms (polar covalent) or even completely transfer b/w atoms (ionic)

$$\Delta EN = 3.0 - 2.1 = 0.9$$

(Polar covalent)



$$\Delta EN = 2.8 - 0.8 = 2.0 \quad (\text{Ionic})$$

7) Explain the significance of the ΔEN values of 0.35 and 1.7.

- Represent transitions b/w non-polar covalent, polar covalent, & ionic bond types.
- ΔEN of 0.35 = 3% ionic character (97% covalent)
- ΔEN of 1.7 = 50% ionic character (50% covalent)
- Used to describe e^- behavior on the bond (shared evenly, unevenly, or transferred)

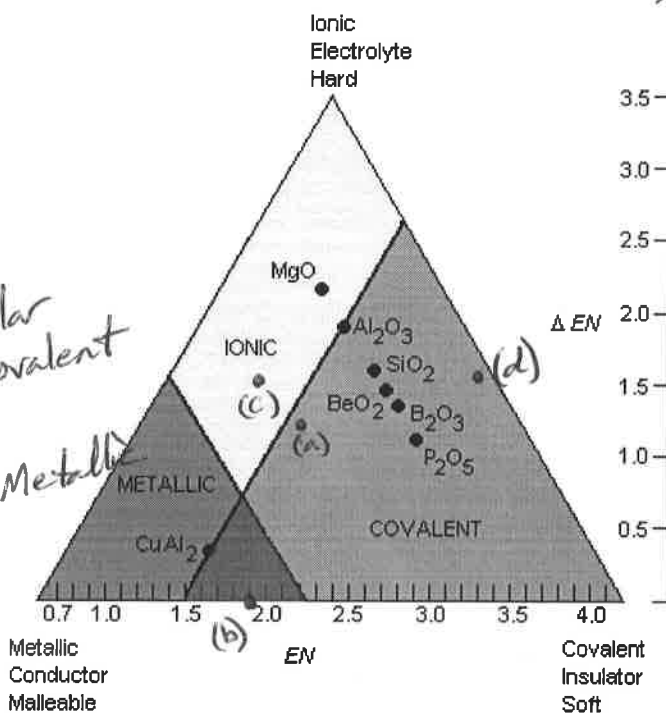
8) Determine the type of bond that would occur in the following compounds according to Dr. Bodner's triangle.

a. GaBr_3
 $\begin{matrix} \rightarrow 2.8 \\ \leftarrow 1.6 \end{matrix}$
 $\Delta EN = 1.2$
 $\overline{EN} = 2.2$
 Polar Covalent

b. NiPb
 $\begin{matrix} \leftarrow 1.9 \\ \rightarrow 1.9 \end{matrix}$
 $\Delta EN = 0$
 $\overline{EN} = 1.9$
 Semi Metallic

c. SrS
 $\begin{matrix} \leftarrow 1.0 \\ \rightarrow 2.5 \end{matrix}$
 $\Delta EN = 1.5$
 $\overline{EN} = 1.75$
 Ionic

d. SF_2
 $\begin{matrix} \leftarrow 2.5 \\ \rightarrow 4.0 \end{matrix}$
 $\Delta EN = 1.5$
 $\overline{EN} = 3.25$
 Polar Covalent



9) List the different types of intramolecular forces. Can you rank them in order of strength? How is their strength determined? If you break intramolecular forces, what type of process is matter undergoing?

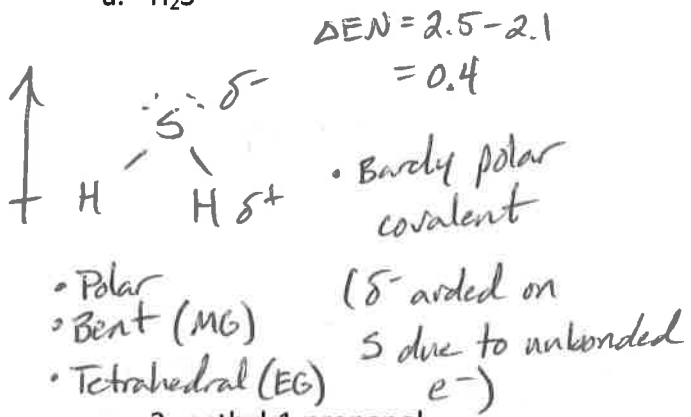
- Non-Polar Covalent, Polar Covalent, Metallic, Ionic
- Can not be ranked by strength (no trend)
- strength measure w/ bond dissociation energy, determined by atomic radius & # of bonds.
- Breaking intra = breaking bonds = chemical reaction

10) What does it mean for a molecule to be polar? What 2 requirements must be met for a molecule to exhibit polarity?

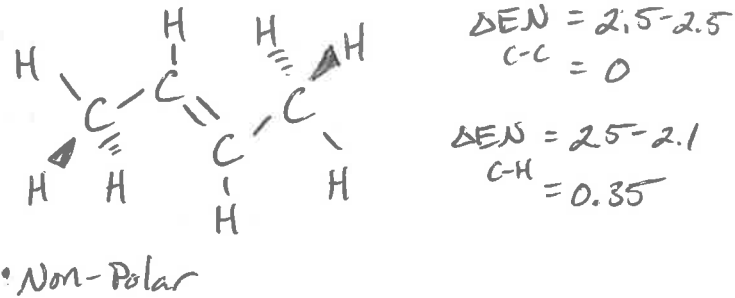
- Opposite regions of + & - on the molecule
- Must be at least 1 polar bond present to form charges & charges must be unevenly distributed (asymmetrical) to form poles.

11) Determine whether or not the following molecules are polar. Include Lewis Dot structures, EN calculations, partial charges, and dipoles.

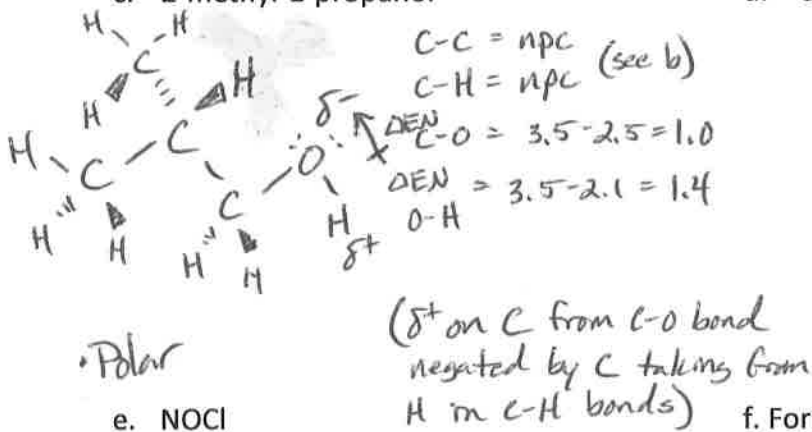
a. H₂S



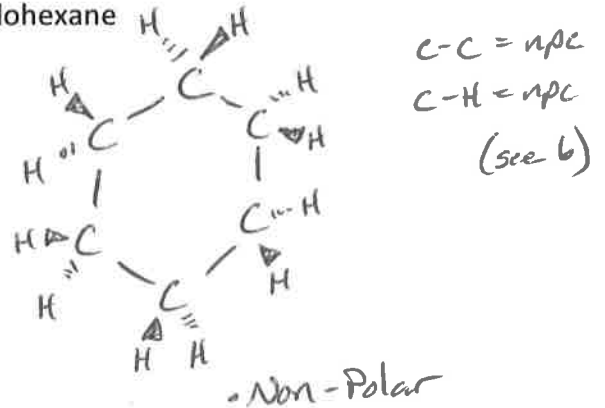
b. 2-Butene



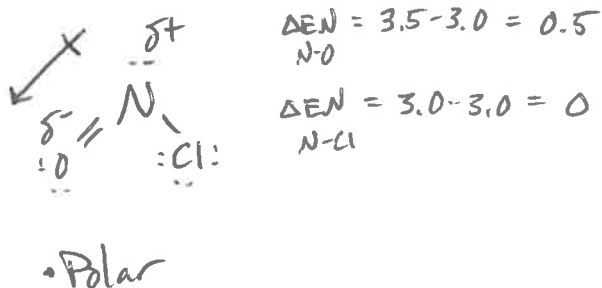
c. 2-methyl-1-propanol



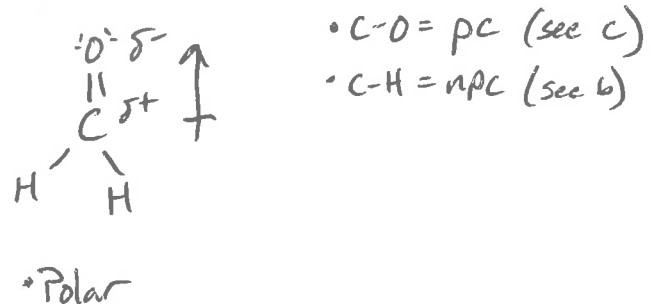
d. Cyclohexane



e. NOCl



f. Formaldehyde (CH₂O)



12) List the intermolecular forces in order of increasing strength and identify the intramolecular force that causes each.

Weakest
↓
Strongest

London Dispersion: weak, momentary δ^+/δ^- caused by uneven e^- distribution that is random & brief. Technically present in all molecules, but is the only IMF a non-polar molecule can exhibit.

Dipole/Dipole: attraction b/w permanent δ^+/δ^- , caused by polar covalent bonds

H-Bonding: strongest dipole/dipole with large δ^+/δ^- (δ^+ on H), caused by polar covalent bonds b/w H and F, O, Cl, N (high EN)

Metalllic: caused by metallic bonds

Ionic: caused by ionic bonds, full charge attraction } Intra & Inter are the same

13) Which would have the stronger intermolecular force? Explain your reasoning for each.

a. CaCl₂ or KBr (Both are ionic IMF)

- Ca & Cl have smaller radii
- Ca & Cl have larger charges (+2 & -1 vs +1 & -1)

• Coulombic Force

b. Ethanol or Ethylaldehyde

- Ethanol: Both are terminal (equal steric hindrance, O-H results in larger δ^+/δ^- than O-C (Hydrogen bonding vs regular dipole/dipole))

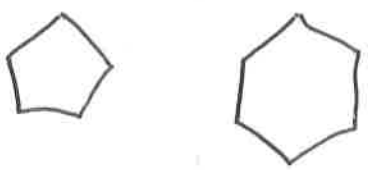
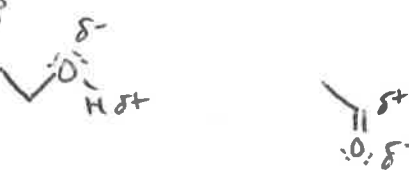
c. Cyclopentane or Cyclohexane

- Cyclohexane: Both non-polar, so London Dispersion would be stronger in the larger molecule.

CaCl₂
higher BP/MP
& lower VP

Ethanol
higher BP/MP
& lower VP

Cyclohexane
higher BP/MP
& lower VP



14) Explain how intermolecular forces account for the following characteristics:

a. Boiling Point Phase changes occur due to changes in the frequency of intact IMFs (solid = high % of IMFs intact restrict movement) (Liquid = fewer IMF, more freedom/fluid motion, but still clings together) (Gases = all IMF, if ideal gas, are not intact, allowing for expansion to fill the container) • Stronger IMFs = \uparrow BP

b. Surface Tension Surface tension, adhesion, cohesion, capillary action are all the result of IMF attraction either b/w molecules or b/w molecule and a surface. Strong IMFs allow for IMF attraction; surface tension is the result of IMF attraction at the surface of a solution/liquid. • Strong surface tension allows for insects to "walk" on water, cause belly flops to hurt, etc.

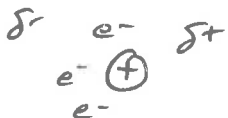
c. Solubility Dissolution (dissolving) is determined by polarity (re IMFs). Like dissolves like: polar dissolves polar using charges to form solvent/solute attraction. Know how IMFs are used to dissolve ionic or covalent compounds, why non-polar/non-polar mix & why non-polar/polar do not.

d. Vapor Pressure Vapor pressure is pressure from a gas that has evaporated/boiled from a solution/liquid. The weaker the IMFs holding the substance in a liquid phase, the lower the boiling point, the more particles that would escape as a gas, increasing vapor pressure.

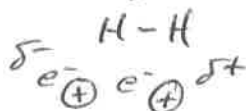
15) Explain the concept of Van der Waals forces and their importance to non-polar molecules.

Van der Waals (London Dispersion) forces are attractions b/w weak, temporary δ^-/δ^+ formed due to uneven e^- distribution.

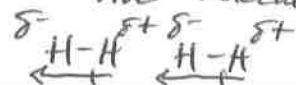
① Random, uneven e^- distribution occurs on an atom



② Causes (Induces) δ^-/δ^+ on bonded neighbor

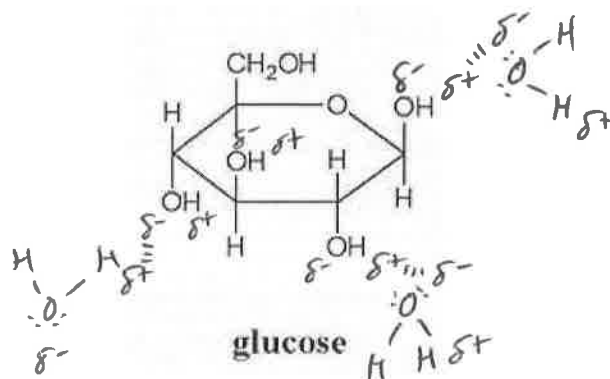


③ Momentary Dipole on a molecule can induce dipoles on other molecules



* Only IMF for non-polar molecules

16) Show how glucose would be dissolved by water at the molecular level. What intermolecular force is responsible for this dissolving process?



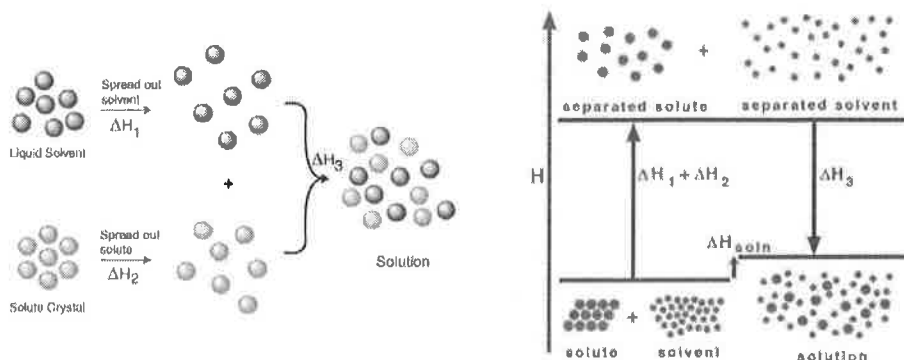
17) Explain the phrase "like dissolves like" by discussing what would be observed after the following action takes place:

100.0mL of butane (MW = 58.12 g/mol, $d = 2.48$ g/mL) and 100.0mL of water (MW = 18.02 g/mol, $d = 1.00$ g/mL) are combined in a beaker and vigorously mixed.

- 1) Vigorous mixing would temporarily intermix the two liquids
- 2) As soon as polar water molecules contact other polar molecules, they will form IMFs (H-bond) and clump together.
- 3) Eventually all of the water will form a common layer by "pushing out" the non-polar butane.
- 4) The butane layer would be on the bottom due to higher density.



18) Describe the diagrams below which depict the process of dissolving a solid solute. Use the concepts of enthalpy and entropy in your explanation.



Dissolving occurs in 2 energetic steps

- 1) IMFs holding solvent to solvent & solute to solute are broken to allow for spreading & dissolution ($+\Delta H$)
- 2) New IMFs are formed b/w solute & solvent ($-\Delta H$)

If more enthalpic energy is used to break IMFs (step 1) than is released during their formation (step 2), the dissolving process will be endothermic. (& visa versa). Dissolving is always entropically favorable.
 ↳ not favorable

19) List the factors that influence the rate of dissolution for a solute and describe why each influences dissolution.

Rate

- 1) Temp/Mixing: Increases particle motion, increasing molecular collisions b/w solvent & solute needed for dissolving.
- 2) solute surface area: Increases # of collisions b/w solvent & solute by having more solute accessible at the surface.
- 3) Concentration: Increased $[]$ means fewer solvent molecules are available to dissolve solute, decreasing the rate of dissolution for those solute particles.
- 4) Pressure: Increases rate of dissolving for gases by forcing gases into the solution at a higher rate.

20) What factors influence the amount of solute one can dissolve in a solution?

1) Temp: Increases solubility of solids (most), decreases solubility of gases.

2) Amount of Solvent: More solvent = More solute can be dissolved.

21) Define the terms saturated, unsaturated, and supersaturated solutions.

Unsaturated: could dissolve more solute at that temp & volume of solvent.

Saturated: has dissolved as much as it can at that temp & volume of solvent.

Super saturated: has dissolved more than it should be able to at that temp and volume of solvent.

22) Use the following solubility curve to determine the type of solution prepared in each scenario.

a. 55g of potassium nitrate is mixed with 50.0mL of water at 60°C.

$$\frac{110 \text{ g KNO}_3}{100 \text{ g H}_2\text{O}} = \frac{x}{50 \text{ g H}_2\text{O}}$$

$x = 55 \text{ g}$
Saturated

b. 50.0g of sodium sulfate is mixed with 90.0mL of water at 10°C.

$$\frac{58 \text{ g Na}_2\text{SO}_4}{100 \text{ g H}_2\text{O}} = \frac{x}{90 \text{ g H}_2\text{O}}$$

$x = 52.2 \text{ g}$
Unsaturated

c. 200.0g of sodium nitrate is mixed with 200.0mL of water at 30°C.

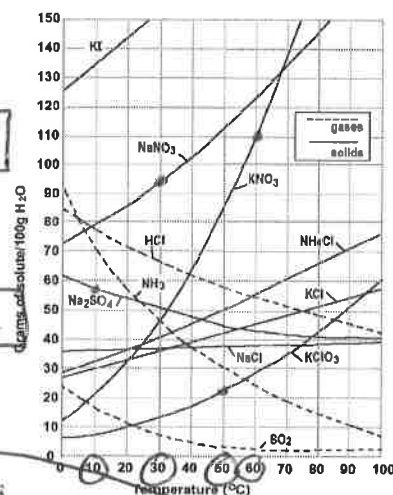
$$\frac{95 \text{ g}}{100 \text{ g H}_2\text{O}} = \frac{x}{200 \text{ g H}_2\text{O}}$$

$x = 190 \text{ g}$
Saturated w/ 10 g not dissolved on the bottom

d. 45g of potassium chlorate is dissolved in 150mL of water at 50°C.

$$\frac{22 \text{ g}}{100 \text{ g H}_2\text{O}} = \frac{x}{150 \text{ g}}$$

$x = 33 \text{ g}$
Supersaturated



23) Explain why increasing temperature increases solubility of most solid solutes, but decreases solubility of gaseous solutes.

↑Temp = ↑ spacing / movement of solvent particles = helps (solids) provide pockets / space for dissolution. Also increase # of collisions per second, aiding dissolving.

↑Temp = same effect on solvent. Makes it easier for gases to (gases) escape the solvent (↑ spacing & ↑ movement = breaks weak IMF attraction of gases)

24) Using Henry's Law, determine the partial pressure of ethanol over 1000.0g of a 95% (by weight) ethanol solution. ($k_H = 111 \text{ L} \cdot \text{atm} / \text{mol}$) (assume $d = 1.0 \text{ g/mL}$)

$$P_{\text{EtOH}} = k_{\text{EtOH}} C_{\text{EtOH}}$$

$$P = \left(111 \frac{\text{L} \cdot \text{atm}}{\text{mol}} \right) \left(\frac{21 \text{ mol}}{1 \text{ L}} \right)$$

$$= 2289 \text{ atm}$$

$$= \boxed{2300 \text{ atm}}$$

$$1000.0 \text{ g} \left(\frac{1 \text{ mL}}{1.0 \text{ g}} \right) \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) = 1.0000 \text{ L}$$

$$1000.0 \text{ g solution} \left(\frac{95 \text{ g EtOH}}{100 \text{ g solution}} \right) = 950.0 \text{ g EtOH}$$

$$950 \text{ g EtOH} \left(\frac{1 \text{ mol EtOH}}{46.07 \text{ g}} \right) = 20.62 \text{ mol}$$

$$= 21 \text{ mol}$$

25) Explain the concept of freezing point depression using IMF.

- Solidification requires the formation of IMFs b/w solvent molecules
- Adding a solute causes them to impair the solvent's ability to form IMF w/ itself, ^(get in the way) making it less likely to form IMFs and therefore freeze. This makes the likelihood of the solvent staying a solution w/ the solute more probable than freezing.

26) An ice storm deposits 55.0kg of ice on your concaved driveway.

- a. Determine the freezing point depression that occurs when a 3.0kg bag of KCl is used to melt the ice. (assume the salt is spread evenly)

$$\Delta T_f = i m K_f$$

$$= 2(0.73 m)(1.86 \frac{^\circ C}{m})$$

$$= \boxed{2.72^\circ C}$$

$$\boxed{\text{new FP} = -2.72^\circ C}$$

$$K_f = 1.86 \frac{^\circ C}{m} \text{ (notes)}$$

$$i = 2 \text{ (K}^+ \text{ \& } \text{Cl}^-)$$

$$m = \frac{\text{mol KCl}}{\text{kg solvent}}$$

$$= \frac{40.24 \text{ mol KCl}}{55.0 \text{ kg H}_2\text{O}}$$

$$= 0.73 \frac{\text{mol}}{\text{kg}}$$

$$\frac{3.0 \text{ kg KCl} \left(\frac{1000 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol}}{74.553 \text{ g}} \right)}{55.0 \text{ kg H}_2\text{O}} = 40.24 \frac{\text{mol KCl}}{\text{kg H}_2\text{O}}$$

- b. Would all of the salt dissolve? Support with math.

Reference solubility chart for KCl, at 0°C ~27g of KCl will dissolve in 100g H₂O

$$\frac{27 \text{ g KCl}}{100 \text{ g H}_2\text{O}} = \frac{x \text{ g}}{55,000 \text{ g H}_2\text{O}}$$

$$x = 14,850 \text{ g could dissolve in the ice}$$

Yes, it is an unsaturated solution holding 3000g of KCl, when >14,000 could dissolve at 0°C.

- c. The salt remains on your driveway until summer, when a rain storm dumps 75.2kg of water, thus dissolving the salt on the driveway. Determine the boiling point elevation that occurs for the rain water. (assume no salt loss and the purity of the rain water)

$$\Delta T_b = i m K_b$$

$$= 2(0.54 m)(0.52 \frac{^\circ C}{m})$$

$$= \boxed{0.56^\circ C}$$

$$\boxed{\text{new BP} = 100.56^\circ C}$$

$$K_b = 0.52 \frac{^\circ C}{m} \text{ (notes)}$$

$$i = 2$$

$$m = \frac{40.24 \text{ mol KCl}}{75.2 \text{ kg H}_2\text{O}}$$

$$= 0.54 m$$

- d. What would be the vapor pressure of water above the salt solution in part c?

N/A (cut out)

- know conceptually it would be less than that of pure water at the same temp ($\uparrow \text{BP} = \downarrow \text{gas}$)
 $= \downarrow \text{VP}$

- 27) Relate the concepts of intramolecular forces, intermolecular forces, boiling point, equilibrium, and vapor pressure.

Intramolecular forces determine the type of intermolecular force a molecule exhibits. The stronger the IMF, the higher the BP of the substance (harder to break IMFs to boil), the higher the BP, the less volatile the substance is, the less/fewer particles escape into the gaseous phase, the lower the vapor pressure.

- 28) What is the osmotic pressure of soda after placing 100.0 mL of a 1.0 M syrup solution (dissolved sugar; $C_6H_{12}O_6$) into 900.0 mL of water at $25^\circ C$?

$$100.0 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{1 \text{ mol } C_6H_{12}O_6}{1 \text{ L}} \right) = 0.100 \text{ mol } C_6H_{12}O_6$$

$$M = \frac{0.100 \text{ mol } C_6H_{12}O_6}{1 \text{ L solution}} = 0.100 \frac{\text{mol}}{\text{L}}$$

$$\pi = iMRT$$

$$= 1 \left(0.100 \frac{\text{mol}}{\text{L}} \right) \left(0.0821 \frac{\text{atm} \cdot \text{L}}{\text{mol} \cdot \text{K}} \right) (273.15 + 25)$$
$$= \boxed{2.4 \text{ atm}}$$

- 29) Go back to #13 and determine which would have a higher vapor pressure and a higher melting point.

