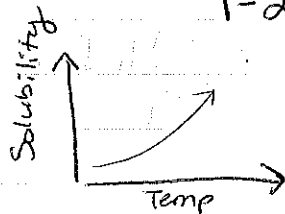


Temp dependence of solvent dissolving:

$$\Delta G = \Delta H - T\Delta S$$

more "—" + of near zero

↳ Increase T
TΔS term gets bigger ⇒ greater solubility!!

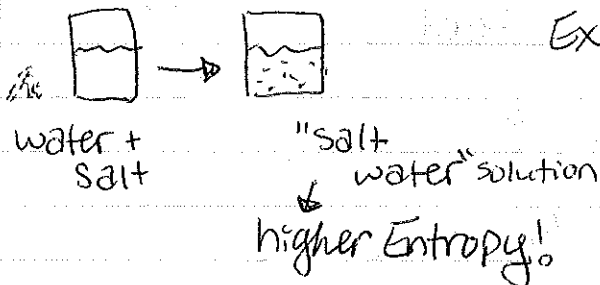


How much dissolves → magnitude of ΔG is important

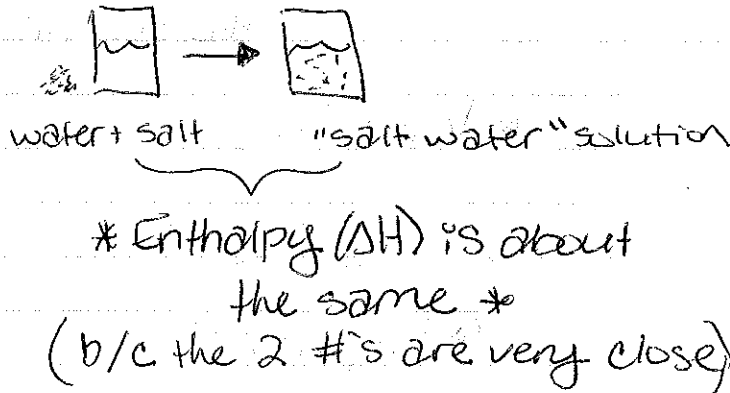
* The bigger negative = greater solubility *

Terms: miscible = capable of being mixed
immiscible = incapable of being mixed

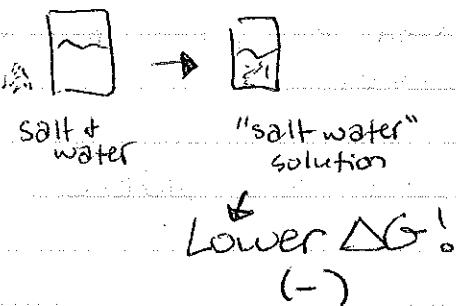
Ex:



Ex:



Ex:



$$\Delta G = \Delta H - T\Delta S$$

$$\Delta G = G_f - G_i$$

↑ must be smaller!

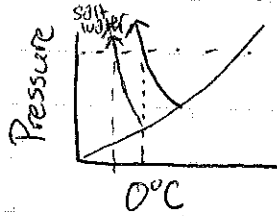
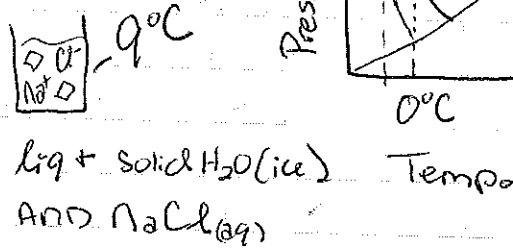
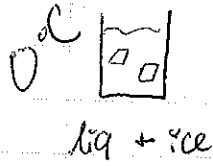
When we make solution

↓
Entropy Increases

↓
Free Energy Decreases

* Therefore more STABLE than pure liquid

- Will freezing Point of a Solution be Different than that of pure solvent?



← Phase Diagram

(So the salt will cause the water to freeze at a lower temp.)

↓
The more you dissolve in, the bigger the entropy, so the bigger the effect

- Vapor Pressure of Solutions

Pure Solvent

↳ higher v.p.

Solution w/ a

nonvolatile solute

↳ lower v.p. b/c free energy is lower

- B.P. Elevation: Solution now more stable

Van't Hoff

factor: • Colligative Properties Depend on Concentration of Solute, Not Type

↑
B.P. elevation,
F.P. depression
v.p. lowering

$i =$ # of ions in formula unit
(for ideal solution)

* Boil. Pt. Elevation (+ change)

$$\Delta T = i K_b m_{\text{solute}}$$

Van't Hoff factor \downarrow K_b constant depends on solvent \leftarrow molality \leftarrow (mol / kg solvent)

* Freezing Pt. Depression (- change)

$$\Delta T = -i K_f m_{\text{solute}}$$

* Change in V.P.

$$\Delta P = -X_{\text{solute}} P^\circ$$

$$P_{\text{solution}} = X_{\text{solvent}} P^\circ$$

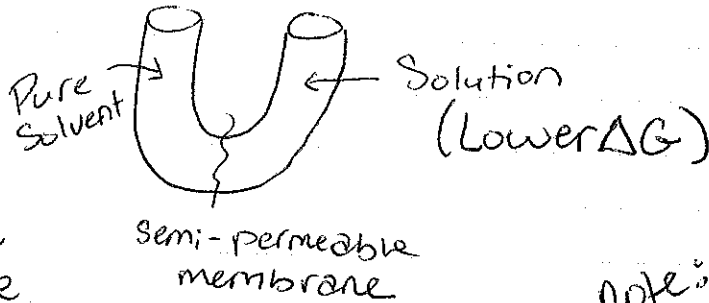
change in V.P. of SOLVENT \leftarrow mol fraction of SOLVENT \leftarrow V.P. of pure SOLVENT

Osmosis

NOTE:

concentration \uparrow
 ΔG

The Solvent will move to the solution side to lower ΔG !!



NOTE: pressure difference depends on concentration of solution !!

* Osmotic Pressure (π) = Pressure needed to stop flow of solvent

$$\pi = i M R T$$

Van't Hoff \uparrow \uparrow molarity \leftarrow gas constant \leftarrow Temp. (K)

(L·atm / mol·K)

Solutions

- Solution = Solvent + Solute
- Solubility: ability for solute to dissolve in solution
- Dissolution: process of dissolving
- 4 types of Solids
 - Metallic \rightarrow all metals ex. Aluminum
 - Ionic \rightarrow network of ionic bonds ex. NaCl
 - Covalent \rightarrow network of covalently-linked (bonds) ex. diamond, graphite
 - Molecular \rightarrow discrete molecule held with IMFs ex. ~~C₁₂H₂₂O₁₁~~ C₁₂H₂₂O₁₁
- Henry's Law
 - $P_{\text{solute}} = K_H X_{\text{solute}}$
 \rightarrow partial pressure
 - As partial pressure of gas solute increase, concentration of dissolved gas decreases
- Like dissolves like \rightarrow by IMFs
- For most mixtures \rightarrow homogeneous solutions:
 - $\Delta H_{\text{soln}} \approx 0$ but positive
 - $\Delta H_{\text{solution}} = \Delta H_{\text{Lattice}} + \Delta H_{\text{solvation}}$
 - IF ΔH is large/positive \Rightarrow salt is insoluble
- Temperature vs. Solvent dissolution for solids
 - For most, as temp \uparrow ; solubility \uparrow
 - For some, as temp \uparrow ; solubility \downarrow
 - ex. $\text{Ce}_2(\text{SO}_4)_3$ or $\text{Na}_2(\text{SO}_4)$
 - because $\Delta H_{\text{solution}} < 0 \rightarrow$ decreases solubility
- Temperature vs. Gas solute dissolution
 - As temp \downarrow ; gas solubility \uparrow
 - $\Delta H_{\text{solution}}$ for a gas $< 0 \rightarrow$ usually
 - $\Delta H_{\text{Lattice}} = 0$ and $\Delta H_{\text{solvation}} < 0$

• Miscible

- Ability for 2 liquids to mix
- Methanol and water
 - $\text{CH}_3\text{OH} \rightarrow$ hydrogen bonding force $>$ dispersion forces
 - i.e. most hydrophilic
- Immiscibility
 - Inability for 2 liquids to mix
 - Oil and water
 - hydrophobic vs. hydrophilic

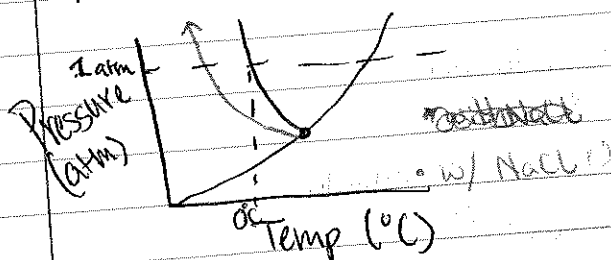
• Water and salt vs. Water and salt solution

- Solution has higher entropy
 - \uparrow in micro states
- Enthalpy are about the same
 - Entropically driven
 - Solvent and solute are similar
- Solution has lower free energy
 - Spontaneous reaction

• Making a Solution

- ΔH close to 0 but positive
- ΔS increasing
- ΔG decreasing, < 0
- Solution is more stable than pure liquid

Phase Diagram

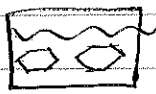


- @ 1 atm, 0°C
both H_2O liquid and ice
have SAME ΔG

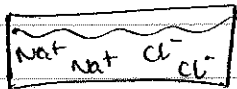
- with salt, point has shifted
 - expanded ΔG to be even lower
 - Freezing point is lower

Boiling Points of Solutions

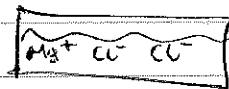
- BP of solution is **higher** than BP of pure water
- Vapor pressure in equilibrium
 - solution @ liquid phase has larger range of stability
- Vapor Pressure of Solutions
 - VP lower for solution b/c ΔG solution is lower
- Sugar does not separate into ions in solvent
 - NaCl does



Sugar
2M ions



Salt
2M ions



MgCl₂
3M ions

↓ lowest freezing point

Colligative Properties (Solution Properties)

- depend on concentration of solute
- ex. VP, FP, BP, etc. and osmosis
- Van't Hoff Factor
 - $i = \#$ ions in solute formula unit
 - Sugar: $i = 1$
 - NaCl: $i = 2$

- Boiling Point Elevation

$$\Delta T = i K_b m_{\text{solute}}$$

K_b depends on solvent \rightarrow BP of solvent
 $m_{\text{solute}} = \text{molality}$

- Freezing Point Depression

$$\Delta T = -i K_f m_{\text{solute}}$$

$K_f = \text{FP of solvent}$

- Change in Vapor Pressure

$$\Delta P = -X_{\text{solute}} P^\circ$$

$X_{\text{solute}} = \text{mole fraction}$

$P^\circ = \text{VP of pure solvent}$

Osmotic Pressure dependent on solute concentration

$$\pi = i M R T$$

Osmotic

\rightarrow molarity