SOLUTIONS

Homogeneous mixture in which one substance is dissolved in another

- **SOLUTE**: substance that is dissolved
- **SOLVENT**: substance doing the dissolving

What factors influence rate of dissolving?

Classification of Solutions

1) **SATURATED**: contains maximum quantity of solute that dissolves at that temperature
2) **UNSATURATED**: contains less than the maximum amount of solute that can be dissolved
3) **SUPERSATURATED**: contains more solute than is possible to be dissolved by warming or evaporating (unstable and temporary)

Classification of Solutions

You NEED to understand what is happening at a molecular level!

Crystallization

Opposite of solution formation

- **Greater** concentration solute particles... **greater** crystallization
- More solute particles bumping into each other and become solid

Aqueous Solution

Solution in which H₂O is the solvent

- Ionic and polar molecules dissolved best... **WHY**?
- **MISCIBLE**: liquids completely mix (alcohol and water)
- **IMMISCIBLE**: don’t mix (oil and water)
Aqueous Solution
Solution in which H₂O is the solvent

• Two Types:
  • ELECTROLYTES: forms ions, conducts electric current (ex: NaCl or MgCl₂) / IONIC
  • NONELECTROLYTES: does not form ions or conduct electric current (ex: Sugar or Ethanol) / POLAR

Solution Process
IMFs
• Interactions in a solution include: solute-solute, solvent-solvent, and solute-solvent
• Solutions will form when the attractive forces between solute and solvent particles are GREATER than those between the solute-solute or solvent-solvent particles alone
  • Ion-Dipole IMFs dominate in water / ionic solutions
  • LDFs dominate in nonpolar solute / nonpolar solvent solutions

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** 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.

Solution Formation and IMFs

**SUMMARY**:

- If energy released when solute-solvent interact is **GREATER** than energy needed to break IMFs in each, then a substance **WILL** dissolve.
- If energy released when solute-solvent interact is **LOWER** than the energy needed to break IMFs in each, then a substance will **NOT** dissolve.

Steps in the Solvation Process:

1. Separation of solute into smaller components (expanding the solute) = ΔH₁
2. Overcoming IMFs between solvent molecules (expanding the solvent) = ΔH₂
3. Formation of solute-solvent IMFs = ΔH₃

\[ ΔH_{\text{soln}} = ΔH₁ + ΔH₂ + ΔH₃ \]

- ΔH₁ = **ENDOthermic**
- ΔH₂ = **ENDOthermic**
- ΔH₃ = **EXOthermic**

Greater OUT than IN... **EXOTHERMIC**!
Greater IN than OUT... **ENDOTHERMIC**!

Examples of **EXOTHERMIC** reactions:

- **Hot Pack** - MgSO₄
  \[ ΔH_{\text{soln}} = -91.2 \text{ kJ/mol} \]
- **Cold Pack** - NH₄NO₃
  \[ ΔH_{\text{soln}} = +26.4 \text{ kJ/mol} \]

If the \( ΔH_{\text{soln}} \) is too highly endothermic, the solution will **NOT** form.
Solutions and Energy Change

How do solutions form when the heat of solution is positive (endothermic)?

Solutions and Entropy

- Spontaneous (no energy added to make the substance dissolve) solutions WILL occur if...
  1) Energy of solution formation is lowered (exothermic)
  2) Some are endothermic but still dissolve due to the spreading out of the components into larger volumes resulting in increased ENTROPY (disorder)

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Solubility Product Constant (K_{sp})

- Saturated solutions have solution in contact with undissolved solute
- To write the K_{sp} expression:
  - ONLY aqueous species (no solids included)
  - UNITLESS
  - Coefficients become exponents in expression
  Ex: \( \text{BaSO}_4(s) \rightleftharpoons \text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) \)
  \( K_{sp} = [\text{Ba}^{2+}] [\text{SO}_4^{2-}] \)
  SMALLER the \( K_{sp} \) the LESS solid will dissolve

Solubility and \( K_{sp} \)

- Solubility and \( K_{sp} \) are DIFFERENT...
- \( K_{sp} \) is unitless
- As \( K_{sp} \) increases so does the amount of solid that dissolves to form a saturated solution
- Solubility measured in g/L or moles/L
  - Amount that dissolves to form a saturated solution
  - Solubility is affected by pH, other ions in solution (common ions), and temperature

If you know the \( K_{sp} \) you can find the solubility of the compound / ions and vice versa!

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*NOTE: 3:1 ratio between calcium ion and solubility of calcium phosphate and 2:1 ratio between phosphate ion and calcium phosphate... this affects problems!

\[ [\text{Ca}^{2+}] \text{ must be 3 times solubility of the compound} \]
\[ [\text{PO}_4^{3-}] \text{ must be 2 times solubility of the compound} \]
Solubility Product Constant ($K_{sp}$)

- **EXAMPLE:**
  
  Calculate the $K_{sp}$ for $\text{Mg}_3(\text{PO}_4)_2$ given that its molar solubility is $3.57 \times 10^{-6}$ M.

  Calculate the **solubility** of $\text{CaF}_2$ and the **concentration** of each ion ($K_{sp} = 4.9 \times 10^{-11}$).

Factors Affecting Solubility

**STRUCTURAL EFFECTS:**

- Polar (hydrophilic) dissolve in polar... more sites of hydrogen bonding = more soluble (ex: water soluble vitamins)
- Nonpolar (hydrophobic) dissolve in nonpolar
- As chains of hydrocarbons become longer, some molecules become less soluble in water and more soluble in nonpolar

Factors Affecting Solubility

**PRESSURE:**

- Little effect on liquids or solids
- Solubility of a gas **INCREASES** with **INCREASING** pressure
- Example: Sodas are bottled under high pressures

Henry’s Law

*Amount of gas dissolved is directly proportional to the pressure of the gas above the solution*

\[ S = kP \]

$S$ = Concentration (usually M) of dissolved gas

$k$ = Constant for solution (mol/L · atm)

$P$ = partial pressure of gaseous solute above solution

**Works best for dilute solutions of gases that don’t dissociate or react with solvent!**
Factors Affecting Solubility

**TEMPERATURE:**
- **SOLIDS...** Increases in temperature cause solids to dissolve more rapidly and usually cause an increase in solubility.
- **GASES...** Solubility of gases decreases with increasing temperature.
- Thermal pollution is bad... higher water temps mean less oxygen dissolved for wildlife!

Factors Affecting Solubility

**TEMPERATURE:**
- Concentration of Solutions
  - Reactions take place when two solutions are mixed.
  - In order to do stoichiometric calculations, the concentration (amount of solute in a solution) must be known.
  - A few ways to express concentration...

Molarity (M)

Moles of solute per liter of solution

\[
M = \frac{\text{Moles of solute}}{\text{Liters of solution}}
\]

**EXAMPLE:**
Calculate the molarity of a solution when 11.5 g of NaOH is dissolved in enough water to make 1.5 liters of solution.

\[
\frac{11.5 \text{ g NaOH}}{1.5 \text{ L soln}} \times \frac{1 \text{ mole NaOH}}{40 \text{ g NaOH}} = 0.19 \text{ M}
\]

Concentration of Solutions

- How many liters of 4.0 M solution can be made using 125 grams of LiBr?

- What mass of oxalic acid, \( \text{H}_2\text{C}_2\text{O}_4 \), is needed to make 250.0 mL of a 0.0500 M solution?

- What is the concentration of a solution that has a volume of 2.5 L and contains 660 g of \( \text{Ca}_3(\text{PO}_4)_2 \)?
**Molality (m)**

Moles of solute per kilogram of solvent

\[ m = \frac{\text{Moles of solute}}{\text{Kilogram of solvent}} \]

- **EXAMPLE:**
  In lab, 1.0 mole of \( \text{C}_2\text{H}_6\text{O}_2 \) is dissolved in 250.0 g of water. Calculate the molality.

\[
\frac{1.0 \text{ mole C}_2\text{H}_6\text{O}_2}{250.0 \text{ g H}_2\text{O}} \times \frac{1 \text{ kg H}_2\text{O}}{1000 \text{ g H}_2\text{O}} = 4.0 \text{ m}
\]

**Mole Fraction**

\# of Moles of one over the total moles

\[ x_a = \frac{n_a}{n_a + n_b + \ldots} \]

- **EXAMPLE:**
  A solution consists of 1.0 g of NaCl mixed with 99.0 g of H\(_2\)O. What is the mole fraction of the solute?

**Mass or Volume %**

\% by Mass or Volume

\[ \frac{\text{g or mL of solute}}{\text{g or mL of solution}} \times 100 \]

- **EXAMPLE:**
  What is the percent concentration of 70.0 g of NaCl dissolved in 350.0 g of water?

\[
\frac{70.0 \text{ g NaCl}}{(70.0 + 350.0) \text{ g soln}} \times 100 = 16.7\%
\]

**Parts Per Million (ppm)**

\[ \text{ppm} = \frac{\text{Grams of solute}}{\text{Grams of solution}} \times 10^6 \]

- **EXAMPLE:**
  A 2.5 g sample of groundwater was found to contain 5.4 x \(10^{-6}\) g of Zn\(^{2+}\). What is the concentration of Zn\(^{2+}\) in parts per million?

\[
\frac{5.4 \times 10^{-6} \text{ g Zn}^{2+}}{2.5 \text{ g soln}} \times 10^6 = 2.2 \text{ ppm}
\]

**Parts Per Billion (ppb)**

\[ \text{ppb} = \frac{\text{Grams of solute}}{\text{Grams of solution}} \times 10^9 \]

- **EXAMPLE:**
  A chemical analysis shows that a water sample contains 8.8 mg of Cd\(^{2+}\) in a 4.00 x \(10^4\) g sample. What is the concentration in parts per billion?

\[
\frac{0.0088 \text{ g Cd}^{2+}}{4.00 \times 10^4 \text{ g soln}} \times 10^9 = 220 \text{ ppb}
\]
Preparing Solutions

- To make a certain concentration of solution, the **SOLID** should be weighed out first and then placed in a volumetric flask.
- Dissolve the solid in **SOME** of the solvent then add the remaining solvent.

**Preparing Solutions**

- To save space, time, and money **PREPARED / CONCENTRATED** solutions are often purchased.
- Water is then added to **DILUTE** the stock solutions to the desired concentration.
- Moles of solute **BEFORE** = Moles of solute **AFTER** (no solute is added during a dilution).
- How much water needs to be added?...

**Preparing Solutions**

- Moles of the solute do **NOT** change after a dilution!

**Dilutions**

\[ M_1 V_1 = M_2 V_2 \]

**EXAMPLE:**

What volume of 16 M H₂SO₄ must be used to prepare 1.5 L of 0.10 M solution?

\[ (16 \text{ M}) (V_1) = (0.10 \text{ M}) (1.5 \text{ L}) \]

\[ V_1 = \frac{0.10 \times 1.5}{16} = 0.0094 \text{ L} \]

or 9.4 mL

Add 1490.6 mL of water to make 0.10 M!

**Dilutions**

• PRACTICE:

How many mL of water should be used to prepare 300.0 mL of 0.750 M NaBr solution using 2.00 M stock solution?

**Solution Stoichiometry**

• Steps for Stoichiometry with **Solution Reactions**:

1) Write and balance the equation
2) Determine solubility
3) Determine the moles of each reactant and find the limiting reagent
4) Find amount of product

Depending on info given, some steps can be skipped!!
Solution Stoichiometry

**EXAMPLES:**

How many grams of Ca(OH)$_2$ are required to react with 350 mL of 0.40 M HCl? Show the balanced equation.

Calculate the mass of solid formed when 1.25 L of 0.050 M lead (II) nitrate and 2.0 L of 0.025 M sodium sulfate are mixed.

Colligative Properties

Depend on the total concentration of solute particles present NOT the type

- IMFs of the solvent are interrupted when the solute is added... changing the properties of the solvent
- Adding solute to the solvent changes the following physical properties: vapor pressure lowering, boiling-point elevation, freezing-point depression, and osmotic pressure

Vapor Pressure Lowering

Presence of a nonvolatile solute LOWERS the VP of a solvent

- Solute gets in the way preventing fewer solvent molecules on the surface from escaping

Vapor Pressure Lowering

(Add salt to the water... higher altitudes would take LONGER to cook because boil at lower temp)
Raoult’s Law

\[ P_A = X_A P^\circ_A \]

IDEAL solutions (like ideal gases) obey this law!

\( P_A \) = Partial pressure exerted by the solvent’s vapor above the solution
\( X_A \) = mole fraction of the solvent
\( P^\circ_A \) = vapor pressure of the pure solvent

**EXAMPLE:**

Glycerin \((C_3H_8O_3)\) is a nonvolatile nonelectrolyte with a density of 1.26 g/mL at 25°C. Calculate the VP at 25°C of a solution made by adding 50.0 mL of glycerin to 500.0 mL of water. The VP of pure water at 25°C is 23.8 torr, and its density is 1.00 g/mL.

Boiling-Point Elevation

Presence of a solute ELEVATES the boiling point

- Nonvolatile solute causes the VP to decrease... more energy must be added to get VP equal to atmospheric
- Solute gets in the way of the solvent molecules and impede their escape from the surface

**ANTIFREEZE** (Operate engine at higher temps)

**COOKING PASTA** (Add salt to the water)

**FRIENDS** keep you COOL
Boiling-Point Elevation

\[ \Delta T_{bp} = K_b m \ i \]

- \( \Delta T_{bp} \) = increase in boiling point (add to normal boiling point for final answer)
- \( K_b \) = proportionality constant that depends on the SOLVENT
- \( m \) = molality
- \( i \) = Van't Hoff factor

Van’t Hoff Factor

- Colligative properties depend on number of particles... soluble **ionic compounds** dissolve to give 2 or more moles of particles (NaCl \( \rightarrow \) Na\(^+\) + Cl\(^-\))
- **Molecular compounds** do not dissociate and form only 1 mole of particles
- Ex: CaCl\(_2\)... \( i = 3 \) / (NH\(_4\))\(_2\)PO\(_4\)... \( i = 4 \)
- In absence of any actual info about \( i \), just use the ideal value \( (i = 1) \)... organic compounds \( i = 1 \)

Boiling-Point Elevation

- **EXAMPLE:**
  At what temperature does a solution of 58.5 g of salt dissolved in 1000.0 g of water boil? \( (K_b = 0.512 \text{ °C/m}; i=2) \)

Freezing-Point Depression

- Presence of a solute LOWERS the freezing point
- Nonvolatile solute causes the VP to decrease... less energy added to get VP of solid and liquid to be same
- Solute gets in the way of the solvent molecules and prevent them from "clustering" together

SURVIVING THE WINTER!

Freezing-Point Depression

- **EXAMPLE:**
  At what temperature does a solution of 58.5 g of salt dissolved in 1000.0 g of water boil? \( (K_b = 0.512 \text{ °C/m}; i=2) \)
Freezing-Point Depression

\[ \Delta T_{fp} = K_f m i \]

- \( \Delta T_{fp} \) = decrease in freezing point (subtract from normal freezing point for final answer)
- \( K_f \) = proportionality constant that depends on the SOLVENT
- \( m \) = molality
- \( i \) = Van't Hoff factor

\[ \Delta T_{fp} = \text{decrease in freezing point} \]
\[ K_f = \text{proportionality constant that depends on the solvent} \]
\[ m = \text{molality} \]
\[ i = \text{Van't Hoff factor} \]

Example:
What is the molar mass of nicotine if 1.921 g dissolved in 48.92 g of water causes it to freeze at -0.450°C. \((K_f = 1.86 \degree C/m)\)

Osmotic Pressure

Pressure that STOPS osmosis

- OSMOSIS... movement of a solvent through selectively permeable membrane toward HIGHER solute conc.

\[ \pi = M R T i \]

- \( \pi \) = osmotic pressure in atm
- \( M \) = molarity of the solution OR \( M = n/V \)
- \( R = 0.082 \text{ L} \cdot \text{atm} / \text{K} \cdot \text{mol} \)
- \( i \) = Van't Hoff factor

Usually used to find MM of unknown... better than FPD or BPE!

Example:
To determine the molar mass of a certain protein, \(1.00 \times 10^{-3}\) g of it was dissolved in enough water to make 1.00 mL of solution. The osmotic pressure of this solution was found to be 1.12 torr at 25.0°C. Calculate the molar mass of the protein.

Phase Diagram
Colloids
Solutions with particles that are too big to dissolve but too small to settle out

- Colloidal dispersions are the dividing line between solutions and heterogeneous mixtures
- Tiny particles suspended in some medium
- **TYNDALL EFFECT:** scattering of light particles... light passes through a solution but is scattered by a colloid

<table>
<thead>
<tr>
<th>Dispersed Particles</th>
<th>Dispersive Medium</th>
<th>Name</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gas</td>
<td>Liquid</td>
<td>Foam</td>
<td>Broken egg white, shaving cream, whipped cream, ice cream soda</td>
</tr>
<tr>
<td>Gas</td>
<td>Solid</td>
<td>Foam</td>
<td>Marshmallow, rubber foam</td>
</tr>
<tr>
<td>Liquid</td>
<td>Gas</td>
<td>Foam</td>
<td>Clouds, fog, mist, hair spray, deodorant spray</td>
</tr>
<tr>
<td>Liquid</td>
<td>Liquid</td>
<td>Emulsion</td>
<td>Milk, mayonnaise, blood</td>
</tr>
<tr>
<td>Liquid</td>
<td>Solid</td>
<td>Gel</td>
<td>Cheese, butter, margarine</td>
</tr>
</tbody>
</table>

Solubility Rules

- **Knowing whether substances are soluble or insoluble tells us if a precipitate forms...**
  - 1) Salts of ammonium and alkali metals are always soluble
  - 2) All chlorides, bromides, and iodides are soluble except when combined with Ag, Hg²⁺, and Pb which are insoluble
  - 3) Chlorates, acetates, and nitrates are soluble
  - 4) Sulfates are soluble except with Ca, Sr, Ba, Hg, Pb, and Ag which are insoluble

- **Practice:**
  Classify each of the following as soluble or insoluble.
  - Sodium carbonate
  - Lead sulfate
  - Barium nitrate
  - Ammonium phosphate
  - Cobalt (II) hydroxide

- **Ex:** Mg(NO₃)₂(aq) + 2 NaOH(aq) → Mg(OH)₂(s) + 2 NaNO₃(aq)