Unit 3:

Chapter 13: Solutions
Homework: Read Chapter 13.
Keep up with assignments

Solutions: Homogeneous mixtures—same throughout
When table salt is mixed with water—the mixture is homogenous
the salt is still there, as you can tell from the taste, or boiling away the water
The component of the solution that changes state or is less is called the solute
The component that keeps its state or is more is called the solvent

Solutions can exist in any phase:
Aqueous NaCl, wine, carbonated water (soda)
Gaseous air
Solid brass metal alloy (Zn/Cu)

Solution Miscibility:
When one substance (solute) dissolves in another (solvent) it is said to be soluble
or miscible. When one substance does not dissolve in another it is said to be
insoluble or immiscible. The solubility of one substance in another depends on two
factors—nature’s tendency toward mixing, and the types of intermolecular
attractive forces

Like dissolves in like: miscible
Polar and nonpolar substances are immiscible
Example 1:
Will C₂H₆ or CH₃OH be more soluble in water?
Will Vitamin C be soluble in fat or water?
Will Vitamin K be soluble in fat or water?
Explain each.

Vitamin C

Vitamin K

Energy Changes and Solution Process:
Break Solute:Solute interactions
Break Solvent:Solvent
Form Solute:Solvent
\[ \Delta H_{\text{soln}} \]

Must break apart like particles and form attractions between solute and solvent particles.

\[ \Delta H_{\text{soln}} \] may be negative, near zero or positive.

Application: hot packs and cold packs
Entropy of solution: \[ \Delta S_{\text{soln}} \] is always positive (more randomness, ways of being)
Example 2:
Which would you predict to have a negative enthalpy of solution and which would have a positive enthalpy of solution? Explain.

Choices: CO\textsubscript{2} gas in water or CaCl\textsubscript{2} solid in water

Concentration (qualitative):
Unsaturated/Saturated/Supersaturated

Concentration (quantitative):

Molarity: M = moles of solute/liters of solution
Mole fraction: X\textsubscript{A} = moles of A/total moles in solution
Mass percent: mass % = mass of solute/mass of total solution \times 100\%
Volume percent: Volume % = volume of solute/volume of solution \times 100\%
Percent mass/volume = mass of solute g/100 ml of solution
Molality: m = moles of solute/kg of solvent
Normality: N = moles of equivalent/liter of solution
    Equivalent = moles of acid or base or electrons transferred.
ppm = mg solute/ kg of solution = mg of solute/ liter of aqueous solution
ppb = \mu g solute/kg of solution

Dilution of Solutions:
M\textsubscript{1}V\textsubscript{1} = M\textsubscript{2}V\textsubscript{2}

Stoichiometry of Solutions:
Use balanced equations, solve for moles: convert mole:mole, solve question.

Example 3:
Solve for the number of grams of sodium acetate and grams of water needed to make 80.0 g of 8.50% NaC\textsubscript{2}H\textsubscript{3}O\textsubscript{2} by mass.
Example 4:
For an aqueous solution of ethanol that is 28.0% C_2H_5OH by mass and whose density is 0.945 g/ml, solve for the solutions concentration in the following units…
a) Molarity
b) molality
c) X_{ethanol}

Example 5:
How do you prepare 250 ml of 0.50 M H_2SO_4 from a stock solution of 6.00 M H_2SO_4?

Example 6:
35.6 ml of 0.282 M NaOH is required to stoichiometrically react with 20.0 ml of phthalic acid H_2C_8H_4O_4 (H_2P) solution. Write the balanced acid-base equation and calculate the Molarity and Normality of the H_2C_8H_4O_4 solution?
pH:
acid-base reactions
$K_w = 1 \times 10^{-14} = [H^+][OH^-]$
pH = - log $[H^+]$
pH < 7: acid
pH = 7: neutral
pH > 7: base

Example 7:
Solve for the pH of a solution of the strong acid 0.020 M HCl.

Factors affecting Solubility:

Temperature effect: solid solute in water

Generally, the solubility of a solid dissolving in water will increase as the temperature increases. Explanation; the $\Delta H_{\text{soln}}$ is often positive, it takes more energy to break attractions in a solid.
**Temperature effect:** gas solute in water

Generally, the solubility of a gas dissolving in water will decrease as the temperature increases.

Come up with the explanation:

![Solubility of Gases in Water at Various Temperatures](image)

**Pressure effect on gas solubility:** Solubility of a specific gas increases as the pressure of that gas above the solution increases.

![Solubility of Gases in Water at Various Pressures (temp = 20°C)](image)
**Henry’s Law:** Solubility = $k P_{\text{solute}}$

<table>
<thead>
<tr>
<th>Gas</th>
<th>$k_H$ at (25°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>O₂</td>
<td>1.3 x 10⁻³ M/atm</td>
</tr>
<tr>
<td>CO₂</td>
<td>3.4 x 10⁻² M/atm</td>
</tr>
<tr>
<td>NH₃</td>
<td>58 M/atm</td>
</tr>
</tbody>
</table>

**Example 8:**
For CO₂ gas in water, solve for the solubility of CO₂ at 25°C and 5.0 atm.

**Colligative Properties:** depends on the number of particles in the solution

- Four properties that depend on the number of particles in the solution:
  - Vapor Pressure Lowering: $P_A = X_A \cdot P^o_A$
  - Freezing Point Depression: $\Delta T_f = i m K_f$
  - Boiling Point Elevation: $\Delta T_b = i m K_b$
  - Osmotic Pressure: $\pi = i M R T$

**van’t Hoff Factor, $i$:** $i = \text{moles of particles/moles of formula units}$

Ionic compounds produce multiple solute particles for each formula unit
The theoretical **van’t Hoff factor, $i$** = the ratio of moles of solute particles to moles of formula units dissolved
The measured van’t Hoff factors are generally less than theoretical due to ion pairing in solution.

<table>
<thead>
<tr>
<th>Van’t Hoff Factors at 0.05 m aqueous solution</th>
<th>solute</th>
<th>$i$ expected</th>
<th>$i$ measured</th>
</tr>
</thead>
<tbody>
<tr>
<td>nonelectrolyte</td>
<td>1</td>
<td>1</td>
<td></td>
</tr>
<tr>
<td>NaCl</td>
<td>2</td>
<td>1.9</td>
<td></td>
</tr>
<tr>
<td>MgCl₂</td>
<td>3</td>
<td>2.7</td>
<td></td>
</tr>
<tr>
<td>FeCl₃</td>
<td>4</td>
<td>3.4</td>
<td></td>
</tr>
<tr>
<td>Weak electrolytes</td>
<td>Number of ions</td>
<td>Just a bit over 1</td>
<td></td>
</tr>
</tbody>
</table>
Example 9:

Predict the “ideal” (maximum) van’t hoff factor, i, for the following aqueous solutions:

a) \( \text{C}_6\text{H}_12\text{O}_6 \text{(aq)} \)

b) \( \text{AlCl}_3 \text{(aq)} \)

c) \( \text{Ca(NO}_3)_2 \text{(aq)} \)

d) \( \text{NaCl} \text{(aq)} \)

e) \( \text{CO}_2 \text{(aq)} \)

f) \( \text{K}_2\text{SO}_4 \text{(aq)} \)

g) \( \text{HF } \text{(aq)} \)

h) \( \text{HCl } \text{(aq)} \)

i) \( \text{HF} \text{ is a weak acid, while HCl is a strong acid. Which acid will deviate the most from the ideal } i? \text{ Explain.} \)

Vapor Pressure Lowering (Raoult’s Law):

For a solution containing a nonvolatile solute…

\[ \Delta P = P^\circ_{\text{solute}} - P_{\text{solution}} = X_{\text{solute}} \bullet P^\circ_{\text{solute}} \]

For an ideal solution when both the solvent and the solute can evaporate, both molecules will be found in the vapor phase

\[ P_{\text{total}} = P_{\text{solute}} + P_{\text{solvent}} \]

The solvent decreases the solute vapor pressure in the same way the solute decreased the solvent’s

\[ P_{\text{solute}} = X_{\text{solute}} \bullet P^\circ_{\text{solute}} \quad \text{and} \quad P_{\text{solvent}} = X_{\text{solvent}} \bullet P^\circ_{\text{solvent}} \]

In ideal solutions, the made solute–solvent interactions are equal to the sum of the broken solute–solute and solvent–solvent interactions

If the solute–solvent interactions are stronger or weaker than the broken interactions the solution is nonideal

When the solute–solvent interactions are stronger than the solute–solute + solvent–solvent, the total vapor pressure of the solution will be less than predicted by Raoult’s Law because the vapor pressures of the solute and solvent are lower than ideal
When the solute–solvent interactions are weaker than the solute–solute + solvent–solvent, the total vapor pressure of the solution will be more than predicted by Raoult’s Law.

**Example 10:**

A solution is prepared by mixing 35.0 grams of acetone \((\text{CH}_3)_2\text{CO}\) and 72.0 grams of chloroform \((\text{CHCl}_3)\) at a constant temperature. The vapor pressure at this temperature of pure acetone is 345 torr, and that of pure chloroform is 295 torr.

a) Assuming ideal behavior, calculate the **mole fractions** in the liquid state of each, the **partial vapor pressure** of each component and the **total vapor pressure** of the solution. (Label each answer clearly)

b) Calculate the **mole fractions** of each component in the **VAPOR** above the liquid. How does this compare to the mole fraction of the liquid? Explain how can this be useful in distillations?
Phase Diagram for a solution compared to the pure solvent:

Freezing Point Depression:
\[ \Delta T_f = i m K_f \]

Example 11:
In making homemade ice cream, the freezing point of the surrounding ice-water salt slush solution is lowered by using 11.5% by mass sodium chloride solution which is observed to freeze at \(-8.02^\circ C\). The normal freezing point of water is exactly 0°C.

a) Calculate the molality of the 11.5% NaCl (aq) solution.
b) Calculate van't Hoff factor, \( i \), in the NaCl solution \( K_f = 1.86^\circ C/m \) for water.
c) Predict the "ideal" (or maximum) van't Hoff factor, \( i \).
d) Solve for the % ionization of NaCl in this solution?
e) Comparing the ideal and calculated van't Hoff factors, would you say this is a strong, weak, or nonelectrolyte?
Boiling Point Elevation:

\[ \Delta T_b = i m K_b \]

Example 12:
A solution was made by dissolving 3.75 g of a nonelectrolyte solute in 104.0 g of acetone. The solution boiled at 56.58°C. The boiling point of pure acetone is 55.95°C, and the \( K_b = 1.71°C/m \). Calculate the molecular weight of the solute.

Osmotic pressure: \( \pi = iMRT \)

The amount of pressure needed to keep osmotic flow from taking place is called the osmotic pressure.

The osmotic pressure, \( \pi \), is directly proportional to the Molarity of the solute particles and Temperature:

\[ \pi = iMRT \quad R = 0.08206 \text{ (atm} \cdot \text{L})/(\text{mol} \cdot \text{K}) \]
Why is it harmful to drink seawater when you are thirsty and stranded in the ocean?
**Example 13:**
A solution is prepared by dissolving 0.137 grams of streptomycin in water and diluting to a volume of 500 ml. This solution has an osmotic pressure of 8.72 torr at 23.6°C. Solve for the molecular weight of streptomycin.

**Uses of Colligative Properties:**
- Fractional Distillation of liquids
- Reverse osmosis water purification
- Antifreeze/coolant in automobile radiators
- Dialysis

**Mixtures:**
- **Solutions** = homogeneous
- **Suspensions** = heterogeneous, separate on standing
  - *Examples of suspensions:* Paint, blood, mud
- **Colloids** = heterogeneous, do not separate on standing
  - *Examples of colloids:* Milk, fog, mayonnaise
    - Mixtures contain particles 2-500nm in diameter that do not separate
    - Small particles are surrounded by charges that repel others in the colloid
    - **Tyndall Effect** – *Colloids scatter light*


<table>
<thead>
<tr>
<th>Classification</th>
<th>Dispersing Substance (Solute-like)</th>
<th>Dispersing Medium (Solvent-like)</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aerosol</td>
<td>Liquid</td>
<td>Gas</td>
<td>Fog (water droplets in air)</td>
</tr>
<tr>
<td>Solid aerosol</td>
<td>Solid</td>
<td>Gas</td>
<td>Smoke (ash in air)</td>
</tr>
<tr>
<td>Foam</td>
<td>Gas</td>
<td>Liquid</td>
<td>Whipped cream (air bubbles in butterfat)</td>
</tr>
<tr>
<td>Emulsion</td>
<td>Liquid</td>
<td>Liquid</td>
<td>Milk (milk fat globules in water)</td>
</tr>
<tr>
<td>Solid emulsion</td>
<td>Liquid</td>
<td>Solid</td>
<td>Opal (water in silica glass)</td>
</tr>
</tbody>
</table>

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Soap:

Triglycerides can be broken down into fatty acid salts and glycerol by treatment with a strong hydroxide solution.

Fatty acid salts have a very polar “head” because it is ionic and a very nonpolar “tail” because it is all C and H.

hydrophilic head and hydrophobic tail

This unique structure allows the fatty acid salts, called soaps, to help oily substances be attracted to water.

micelle formation
emulsification

![Image of sodium stearate](source)

![Image of micelle](source)
**Solutions: Example Problems:**

1. The solubility of KNO₃ is 45.0 g of KNO₃ per 100 grams of water at 30°C.
   a) What is the mass % of KNO₃ in the above-saturated solution at 30°C?

   b) If a solution contains 12.0 g or KNO₃ in 40.0 g of water is this solution unsaturated, saturated, or supersaturated? (compare with part a, the saturated solution)

2. An aqueous oven-cleaning solution is 40.0% (by mass) NaOH. If one jar of this product contains 1.00 kg of solution, how many moles of solid NaOH are dissolved in the solution?

3. a. What volume (ml) of 12.0 M HCl is needed to prepare 300. ml of 1.90 M HCl?

   b. Briefly explain how you would accurately prepare this solution in a lab environment.
4. A 100.0 gram bottle contains 20.0 % C_{12}H_{22}O_{11} (sugar) solution by mass. What is its concentration in units of Molarity if the density is 1.08 g/ml? (C = 12.0 g/mol; H = 1.008 g/mol; O = 16.0 g/mol)

a) Solve for grams of sugar you have in 100 grams of solution

b) Solve for the moles of sugar

c) Solve for the total volume of solution

d) Solve for molarity of the solution

5. Balance the equation for the titration problem below.

\[ \text{H}_2\text{C}_2\text{O}_4 \text{(aq)} \ + \ \text{NaOH} \text{(aq)} \ \rightarrow \ \text{H}_2\text{O} \text{(aq)} \ + \ \text{Na}_2\text{C}_2\text{O}_4 \text{(aq)} \]

A chemist pipettes 10.0 ml of \( \text{H}_2\text{C}_2\text{O}_4 \) solution into an Erlenmeyer flask and adds some phenolphthalein solution as an indicator. Standardized 0.262 M sodium hydroxide was placed into the buret and a titration reaction is performed until the solution stays pink. Using the titration data below, calculate the molarity of the acetic acid.

**NaOH buret readings:**

- final 29.5 ml
- initial 0.0 ml
6. Vitamin C has the chemical name ascorbic acid. It can be abbreviated as HAsc. Given that 0.880 grams of HAsc can be neutralized with 33.55 ml of 0.150 M NaOH, solve for the molecular weight of Vitamin C. The balanced reaction is:

\[
\text{HAsc (aq) + NaOH (aq) } \rightarrow \text{NaAsc (aq) + H}_2\text{O (l)}
\]

a) Solve for moles of NaOH used in the titration

b) Solve for moles of ascorbic acid

c) Solve for molecular weight of the ascorbic acid

7. Two shipwreck survivors were rescued from a life raft. One had drunk seawater, while the other had not. The one who had drunk seawater was more severely dehydrated that the one who did not. Explain.

8. What is the Tyndall effect and what will it test for?

9. What will be the melting/freezing point of brass made of 10.0% Zn and 90.0% Cu by mass? The melting/freezing point constant, \(K_f\), for copper is 23°C/m. Pure copper melts at 1083°C. the van't Hoff factor, \(i\), is 1.

10. Three beakers contain 0.080\text{m} of the following aqueous, soluble solutions Ca(NO_3)_2, C_{12}H_{22}O_{11}, Al(NO_3)_3, and KBr.

\(a\) Would you expect the solutions to boil below or above the normal boiling point of water, 100°C?

\(b\) Place the aqueous 0.080\text{m} solutions above in increasing order starting with the solution that should have the lowest normal boiling point and end with the solution having the highest normal boiling point. Explain your reasoning. These are dilute solutions with colligative properties and not pure substances.

11. What is the equivalent weight of H_2SO_4?

12. Calculate the molarity of a solution of H_2SO_4 if 32.0 ml of the solution is required to neutralize 0.228 g of KOH.
13. There is 225 ml of a solution that is 19.6% H₂SO₄ by mass. The density of the solution is 1.14 g/ml.
   a) How many grams of H₂SO₄ are in the solution?
   b) What is the molality of the solution?
   c) What are the mole fractions of H₂SO₄ and H₂O in the solution?
   d) What is the Molarity of the solution?
   e) What is the Normality of the solution?

14. You are a technician in a chemical company. There are 40 titrations you must finish before you go home and each needs about 30 ml of a 0.200 M Na₂CO₃ solution.
   a) What volume of this sodium carbonate solution do you need to prepare?
   b) How do you prepare the solution if you have H₂O and solid Na₂CO₃ available?
   c) How do you prepare the solution if you have H₂O and 2.50 M Na₂CO₃ solution?

15. The heat of solution ΔHₜₜₜ depend on solute-solute attractions, solvent-solvent attractions, and solute-solvent attractions.
   a) Which attractions must be separated apart (is this usually endothermic or exothermic) and which attractions are made (is this usually endothermic or exothermic)?
   b) Why are most solids more soluble in warm water while gases are more soluble in cold water?

16. Given a 0.010 M Al(NO₃)₃ solution and a 0.010 M C₆H₁₂O₆ solution at 25°C
   a) Which would have the larger boiling point elevation? Explain.
   b) Which would have the greater deviation from its "ideal" van't Hoff factor, i. Explain.

17. Calculate the approximate molecular weight of an unknown protein if a 0.239 gram of this protein dissolved in 79.3 ml of H₂O, produces a solution that has an osmotic pressure of 17.34 torr at 25.0°C.

18. Fill in with the words dilute or concentrated:
   If 2 liquids are separated by a semipermeable membrane, the more _________ solution will always flow to the more ________________ solution at a faster rate. This causes the more ________________ side to rise up higher. Osmotic pressure is the amount of external pressure required on the more ________________ side to equalize the 2 levels.