Chapter 14

Physical Properties of Solutions

Concentration Units

**Molarity (M)** = moles solute / Liters of solution

**Percent by Mass (weight)**

**Percent by Volume**

**Parts per million** - (for dilute aqueous solutions 1ppm = 1mg/L, since the density is close to that of water at 1.00g/mL)

**Parts per billion** - (1ppb = 1µg/L)

**Parts per trillion** - (1ppt = 1ng/L)

These units can be in terms of numbers of particles or mass (usually mass for liquids, particles for gases)

Example:

1. Show that for dilute solutions, 1ppm is approximately equal to 1mg/L

Result:

At 1.00g/mL for water there are 1,000g/L

In one milligram there is .001g

.001g / 1,000g = 1x10⁻⁶ = 1 / 1,000,000 = 1ppm

Example:

2. What is the mass percent sucrose (table sugar) in the solution obtained by mixing 225g of an aqueous solution that is 6.25% sucrose by mass with 135g of an aqueous solution that is 8.20% sucrose by mass?

---

**TABLE 14.1 Some Common Solutions**

<table>
<thead>
<tr>
<th>Solution</th>
<th>Components</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gaseous solutions</td>
<td>N₂, O₂, and several others</td>
</tr>
<tr>
<td>Natural gas</td>
<td>CH₄, C₂H₆, and several others</td>
</tr>
<tr>
<td>Liquid solutions</td>
<td>H₂O, NaCl, and many others</td>
</tr>
<tr>
<td>Seawater</td>
<td>mL of H₂O, HC₂H₃O₂ (acetic acid)</td>
</tr>
<tr>
<td>Soda pop</td>
<td>H₂O, CO₂, C₆H₅O₂⁻ (sucrose), and several others</td>
</tr>
<tr>
<td>Yellow brass</td>
<td>Cu, Zn</td>
</tr>
<tr>
<td>Palladium-hydrogen</td>
<td>Pd, H₂</td>
</tr>
</tbody>
</table>
Example:
2. What is the mass percent sucrose (table sugar) in the solution obtained by mixing 225g of an aqueous solution that is 6.25% sucrose by mass with 135g of an aqueous solution that is 8.20% sucrose by mass?

Result:
\[
\frac{(225g)(.0625) + (135g)(.0820)}{(225g + 135g)} = .0698 = 6.98\
\]

Example:
3.
   a. Assume that the volumes are additive, and determine the volume percent toluene \( \text{C}_6\text{H}_5\text{CH}_3 \), in a solution made by mixing 40.0mL of toluene with 75.0mL of benzene, \( \text{C}_6\text{H}_6 \)

   b. At 20°C, the density of benzene is 0.879g/mL and that of toluene is 0.866g/mL. For the solution described in 3a, what is the mass percentage of toluene and the resulting solution density?

Result:
   a. \( \frac{40.0\text{mL}}{40.0\text{mL} + 75.0\text{mL}} = .3478 = 34.8\% \)

   b. Mass\% = \( \frac{(40.0\text{mL})(0.866\text{g/mL})}{(40.0\text{mL})(0.866\text{g/mL}) + (75.0\text{mL})(0.879\text{g/mL})} \) = .344 = 34.4%

   Density = \( \frac{(40.0\text{mL})(0.866\text{g/mL}) + (75.0\text{mL})(0.879\text{g/mL})}{40.0\text{mL} + 75.0\text{mL}} \) = .874g/mL

Example:
Calculate the molality of a sulfuric acid solution containing 24.4g of sulfuric acid in 198g of water. The molar mass of sulfuric acid is 98.08g/mol.

Result:
\[
m = \frac{(24.4g/98.08g/mol)}{(.198kg)} = 1.28\text{molal}
\]
Example:
1. The density of a 2.45M aqueous solution of methanol is 0.976g/mL. What is the molality of the solution? The molar mass of methanol is 32.04g/mol.
2. Calculate the molality of a 44.6% (by mass) aqueous solution of sodium chloride.

Results:
1. \[
\frac{2.45\text{mol}}{((1,000\text{mL})(0.976\text{g/mL}) - (2.45\text{mol})(32.04\text{g/mol}))/1,000\text{g/kg}} = 2.73\text{m}
\]
2. In 100.g of solution there are 44.6g of NaCl
\[
(44.6g/58.44g/mol) / [(100g – 44.6g)/1,000g/kg] = 13.8\text{m}
\]

Mole Fraction (\(x_i\)) and Mole Percent

\(x_i = \frac{\text{moles of component } i}{\text{total moles of solution}}\)

To convert to percent, multiply by 100

Example:
If 10.5g of sodium chloride and 12.6g of potassium chloride are dissolved in 250.g of water, what is the mole fraction of the chloride ion?

Result:

\[
\begin{align*}
\text{Mol NaCl} &= 10.5g / 58.44g/mol = .17967\text{mol NaCl} \\
&\quad / (1.7967\text{mol Na}^+, .17967\text{mol Cl}^-) \\
\text{Mol KCl} &= 12.6g / 74.55g/mol = .16901\text{mol KCl} (.16901\text{mol K}^+, .16901\text{mol Cl}^-) \\
\text{Mol H}_2\text{O} &= 250.g / 18.016g/mol = 13.877\text{mol H}_2\text{O} \\
\text{Mol fraction} &= (1.7967 + .16901) / (1.7967(2) + .16901(2) + 13.877) \\
&= .0239 \quad \text{(or 2.39mol %)}
\end{align*}
\]

Example:
The density of acetonitrile, CH\(_3\)CN, is 0.786g/mL. and the density of methanol, CH\(_3\)OH, is 0.791g/mL. A solution is made by dissolving 20.0mL CH\(_3\)OH in 100.0mL of CH\(_3\)CN.

a. What is the mole fraction of methanol in the solution?

b. What is the molality of the solution?

c. Assuming that the volumes are additive, what is the molarity of CH\(_3\)OH in the solution?

Example:
The density of acetonitrile, CH\(_3\)CN, is 0.786g/mL, and the density of methanol, CH\(_3\)OH, is 0.791g/mL. A solution is made by dissolving 20.0mL CH\(_3\)OH in 100.0mL of CH\(_3\)CN.

a. What is the mol fraction of methanol in the solution?

b. What is the molality of the solution?

c. Assuming that the volumes are additive, what is the molarity of CH\(_3\)OH in the solution?

Answers:

a. Moles CH\(_3\)OH = 20.0mL(0.791g/mL)(1mol/32.04g) = .49376mol
Moles CH\(_3\)CN = 100.0mL(0.786g/mL)(1mol/41.05g) = 1.91474mol
mole fraction CH\(_3\)OH = (0.49376)/(0.49376+1.91474) = 0.205

b. Molality = .49376mol CH\(_3\)OH / 0.0786kg CH\(_3\)CN = 6.28m

c. Molarity = .49376mol CH\(_3\)OH / 0.120L solution = 4.11M
Physical Properties of Solutions

Solution Energy

There are four possibilities in solution formation:

1. All intermolecular forces are of comparable strength in which case: \( \Delta H_{\text{soln}} = 0 \). (We also expect that \( \Delta V_{\text{soln}} = 0 \))
   This type of solution is known as an ideal solution.

2. Intermolecular forces between solute and solvent molecules are stronger than the separate intermolecular forces.
   \( \Delta H_{\text{soln}} < 0 \) exothermic
   and \( \Delta V_{\text{soln}} < 0 \)
   This is a nonideal solution.

3. Intermolecular forces between solute and solvent molecules are somewhat weaker than the separate intermolecular forces.
   \( \Delta H_{\text{soln}} > 0 \) endothermic
   A solution may or may not form depending on other conditions (entropy) and is considered a nonideal solution.

4. Intermolecular forces between solute and solvent molecules are much weaker than the separate intermolecular forces.
   \( \Delta H_{\text{soln}} \gg 0 \) endothermic
   The enthalpy of the solution is so positive that a solution will not form and instead creates a heterogeneous mixture.

**Special Note:** As a general rule, remember like dissolves like, meaning that polar substances tend to dissolve in polar substances and nonpolar substances dissolve in nonpolar substances.

Image: The sum of two 50.0mL volumes of water and ethyl alcohol is less than 100.0mL.

Question: Identify the substances and arrange them in the expected order of increasing solubility in water.
Based on molecular length and polarity:

\[ c \text{ (hexane)} < b \text{ (hexanol)} < d \text{ (butanoic acid)} < a \text{ (ethanoic (acetic) acid)} \]

**Aqueous Solutions of Ionic Compounds**

Ions dissolved in aqueous solution have a certain number of water molecules surrounding them. They are said to be hydrated.

The solubility of an ionic compound is determined largely by the competition between interionic attractions that hold ions in a crystal and the ion-dipole attractions that pull them into solution.

It is difficult to predict the solubility of some ionic compounds.

**Heat of Solution**

Heat related to the solvation process with respect to ionic compounds in water.

Energy must be supplied to separate the ions in the lattice against their attractive forces (-\( \Delta H_{\text{lattice}} \)).

Energy is evolved when the individual ions are transferred into water (\( \Delta H_{\text{hydration}} \)).

Overall: \( \Delta H_{\text{soln}} = -\Delta H_{\text{lattice}} + \Delta H_{\text{hydration}} \)

General Rule: To be soluble, salts must have \( \Delta H_{\text{soln}} \) that is exothermic or only slightly endothermic.

(Enthalpy and entropy (chp 19) must be taken into account)

Note that the energy quantities are affected by ion sizes and charges.

**Question:**

Given \( \Delta H^\circ_{\text{f}}(\text{s}) = -425.9 \text{kJ/mol} \) and \( \Delta H^\circ_{\text{f}}(\text{aq, 1m}) = -469.2 \text{kJ/mol} \) for sodium hydroxide, calculate the enthalpy of a solution of NaOH.

**Answer:**

\[
425.9 \text{kJ/mol} + (-469.2 \text{kJ/mol}) = -43.3 \text{kJ/mol} \text{ (exothermic)}
\]

**Equilibrium in Solution Formation**

Liquids that mix in all proportions are said to be miscible.

Dynamic equilibrium of solution concentration occurs when the rate of crystallization = rate of ions leaving the crystal. The solution is said to be saturated.

95% of ionic compounds have aqueous solubilities that increase with temperature. Some change very little and an even smaller portion becomes less soluble with increased temperature.

Solubility is expressed on a solubility curve which plots solubility in g solute per 100 g water, as a function of temperature. (p672)
Carefully cooling a saturated solution can lead to *supersaturation*, a condition in which more solute is dissolved in solution than is normally possible. The addition of a seed crystal causes the oversaturated solute to quickly crystallize.

**Sodium Acetate**

Supersaturated solution with a "seed crystal" dropped in causes immediate crystallization.

---

**Gas Solubility**

Most gases become less soluble with temperature. **External pressure** has very little effect on the solubility of solids and liquids but *increases the solubility of gases with increased pressure*.

At constant temperature, the solubility \( S_g \) of a gas is directly proportional to the partial pressure of the gas \( P_g \) in equilibrium with the solution.

**Henry’s Law**

\[
S_g = k_H P_g
\]

\( k_H \) = constant (at a given temperature) that is dependent upon the gas and solvent (in mg/g atm, M/mmHg, etc., i.e. some concentration unit over pressure)

<table>
<thead>
<tr>
<th>Gas</th>
<th>( k_H ) (mg/mmHg)</th>
</tr>
</thead>
<tbody>
<tr>
<td>( N_2 )</td>
<td>( 8.42 \times 10^{-7} )</td>
</tr>
<tr>
<td>( O_2 )</td>
<td>( 1.45 \times 10^{-4} )</td>
</tr>
<tr>
<td>( CO_2 )</td>
<td>( 4.48 \times 10^{-5} )</td>
</tr>
</tbody>
</table>

Effect of gas pressure on aqueous solubilities of gases (at 20°C)

Henry’s Law

\[
S_g = k_H P_g
\]
Effect of temperature on solubility of gases

Oxygen only makes up about 23% of air by mass but constitutes about 35% of the air dissolved in water, due to greater solubility.

Question:
At 25°C and 1atm gas pressure, the solubility of CO₂(g) is 149mg/100g water. When air at 25°C and 1 atm is in equilibrium with water, what is the concentration of dissolved CO₂, in mg/100g water? Air contains 0.037 mole% CO₂(g).

This question is an application of Henry's Law:

\[ k = \frac{S}{P_{\text{gas}}} = \frac{(149 \text{mg CO}_2/100 \text{g H}_2\text{O})}{1 \text{atm}} = 1.49 \text{mg CO}_2/\text{g water atm} \]

\[ n(\text{CO}_2)/n(\text{Air}) = P(\text{CO}_2)/P(\text{air}) \]

\[ 0.00037 = P(\text{CO}_2)/1 \text{atm} \]

(From gas laws: Law of Partial Pressure)

\[ P(\text{CO}_2) = 0.00037 \text{atm} \]

\[ S = P_{\text{gas}}k = (0.00037)(1.49) = 5.5 \times 10^{-4} \text{mg CO}_2/\text{g H}_2\text{O} \]

\[ = 5.5 \times 10^{-2} \text{mg CO}_2/100 \text{g H}_2\text{O} \]

Question:
How many milligrams of a mixture containing equal numbers of moles of CH₄ and N₂ at 10.0atm total pressure will dissolve in 1.00L of water at 20°C (d_H₂O = 1.00g/mL)?

Answer:

\[ X_{\text{N}_2} = P_{\text{N}_2}/P_{\text{mixture}} \]

\[ 0.50 = P_{\text{N}_2}/10 \text{atm} \]

\[ P_{\text{N}_2} = 5.0 \text{atm} \]

\[ X_{\text{CH}_4} = P_{\text{CH}_4}/P_{\text{mixture}} \]

\[ 0.50 = P_{\text{CH}_4}/10 \text{atm} \]

\[ P_{\text{CH}_4} = 5.0 \text{atm} \]

From the solubility graph:
At 5.0atm N₂ is 9.0mg gas/100g H₂O
At 5.0atm CH₄ is 12.0mg gas/100g H₂O

\[ \text{?mass N}_2 = (9.0 \text{mg} / 100 \text{g H}_2\text{O}) \times (1 \text{gH}_2\text{O}/1 \text{mL}) \times (1 \text{mL}/10^{-3} \text{L}) \times 1.00 \text{L} = 90.0 \text{mg N}_2 \]

\[ \text{?mass CH}_4 = (12.0 \text{mg} / 100 \text{g H}_2\text{O}) \times (1 \text{gH}_2\text{O}/1 \text{mL}) \times (1 \text{mL}/10^{-3} \text{L}) \times 1.00 \text{L} = 120 \text{mg CH}_4 \]

Total mass = 90.0 mg N₂ + 120mg CH₄ = \(210 \times 10^3\)mg mixture.

Question:
What is the concentration of O₂ in a fresh water stream in equilibrium with air at 25°C and 1.0atm? The mole fraction of O₂ in air is 0.21 and kₕ for O₂ at this temperature is \(1.66 \times 10^{-6} \text{M/mmHg}\). Express the answer in grams of O₂ per liter of solution.
What is the concentration of O\textsubscript{2} in a fresh water stream in equilibrium with air at 25°C and 1.0atm? The mole fraction of O\textsubscript{2} in air is 0.21 and \( k_H \) for O\textsubscript{2} at this temperature is 1.66x10^{-6} M/mmHg. Express the answer in grams of O\textsubscript{2} per liter of solution.

**Answer:**
The mole fraction is proportional to the pressure fraction

\[ P_{O_2} = (1.0 \text{atm})(760 \text{mmHg}/1 \text{atm})(0.21) = 159.6 = 160 \text{mmHg} \]

Solubility = (1.66x10^{-6} mol/L mmHg)(160 mmHg) = 2.7x10^{-4} mol/L

(2.7x10^{-4} mol/L)(32.0 g/mol) = 0.0085 g/L

---

**Physical Properties of Solutions**

**Colligative Properties**

**Colligative Properties:** Physical properties of solutions that depend on the number of solute particles present but not on the identity of the solute.

Types:
- Vapor Pressure Lowering
- Freezing Point Depression
- Boiling Point Elevation
- Osmotic Pressure

Discussion will begin with solutes and solvents in molecular form (nonelectrolytes)

**Changes in Vapor Pressure: Raoult’s Law**
The vapor pressure of the solvent over the solution is lower than the vapor pressure of the pure solvent.
The vapor pressure of the solvent, \( P_{\text{solute}} \), is proportional to the relative number of solvent molecules in the solution; that is, the solvent vapor pressure is proportional to the solvent mole fraction, \( P_{\text{solute}} = X_{\text{solute}} \)

\[ P_{\text{solute}} = X_{\text{solute}} P_{\text{solvent}} \]

(Where \( P \) is the normal vapor pressure of the solvent and \( X \) is the mole fraction of the solvent)

Alternatively,

\[ \Delta P_{\text{solute}} = -X_{\text{solute}} P_{\text{solvent}} \]

**Example:**
Assume you dissolve 10.0 g of sucrose (C\textsubscript{12}H\textsubscript{22}O\textsubscript{11}) in 225 mL (225 g) of water and warm the water to 60.0°C. What is the vapor pressure of the water over this solution? (The vapor pressure of water at 60.0°C is 149.4 torr)

**Answer:**

(10.0 g sucrose / 342.2965 g/mol) = 0.0292 mol

(225 g water / 18.016 g/mol) = 12.489 mol

\[ P = X_{\text{solute}} P_{\text{solvent}} \]

\[ P = (0.0292 + 12.489) \times 149.4 \text{torr} = 149.0515 = 149.1 \text{torr} \]

**Consequences of Raoult’s Law:**

**Boiling Point Elevation**

\[ \Delta T_{bp} = K_{bp} \text{m}_{\text{solute}} \]

\( K_{bp} \) = molal boiling point elevation constant (°C/m)

Where the constant is based on the solvent.
Since the vapor pressure lowering lowers the entire vapor pressure curve, freezing point depression also occurs:

\[ \Delta T_{fp} = K_{fp} \text{mol solute} \]

\( K_{fp} \) is the freezing point depression constant (°C/m)

Questions:

1. In the northern United States, summer cottages are usually closed up for the winter. When doing so, the owners “winterize” the plumbing by putting antifreeze in the toilet tanks, for example. Will adding 525g of HOCH\(_2\)CH\(_2\)OH (ethylene glycol) to 3.00kg of water ensure that the water will not freeze at -25°C? (K\(_{fp}\) for H\(_2\)O is -1.86°C/m)

Answer:

The molar mass of ethylene glycol is 62.0678g/mol.

moles = 525g / 62.0678g/mol = 8.45849mol

\[ \Delta T = K_{fp} \text{mol} \]

\[ \Delta T = (-1.86°C/m)(8.45849mol) = -5.24°C \]

This will not be enough protection for the home.

2. What quantity of ethylene glycol, HOCH\(_2\)CH\(_2\)OH, must be added to 125g of water to raise the boiling point by 1.0°C? Express your answer in grams. (K\(_{bp}\) = 0.5121 °C/m)

Answer:

\[ \Delta T = K_{bp} \text{mol} \]

\[ 1.0°C = (0.5121°C/m)(\text{mol EG}/125kg) \]

Mol EG = 0.24409

0.24409mol (62.0678g/mol) = 15.15g = 15g
3. Crystals of the beautiful blue hydrocarbon, azulene (mass, 0.640g), which has an empirical formula of $C_5H_4$, are dissolved in 99.0g of benzene. The boiling point of the solution is 80.23°C. What is the molecular formula of azulene? 

$K_{bp} = 2.53 \ \degree C/m$, normal boiling point $= 80.10 \degree C$

Answer:

$\Delta T = 80.23 \degree C - 80.10 \degree C = .13 \degree C$

$.13 \degree C = (+2.53 \degree C/m)(mol azulene/.0990kg)$

Mol azulene = .005087mol

Molar mass = g/mol = 0.640g / .005087mol = 125.8g/mol

Empirical mass for $C_5H_4$ = 64.0853g/mol

Molecular formula is $C_{10}H_{8}$

The van’t Hoff factor:

Accounts for the production of ions in solution

$\Delta T_{fp} (measured) = K_{fp}(m)(i)$

Where $i$ is called the van’t Hoff factor. (See table 14.4, page 680)

Table 14.4 Freezing Point Depressions of Some Ionic Solutions

<table>
<thead>
<tr>
<th>Mass %</th>
<th>m (mol/kg)</th>
<th>$\Delta T_{p}$ (measured, °C)</th>
<th>$\Delta T_{p}$ (calculated, °C)</th>
<th>$\Delta T_{p}$, measured</th>
<th>$\Delta T_{p}$, calculated</th>
<th>van’t Hoff factor</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.00</td>
<td>0.0020</td>
<td>-0.0433</td>
<td>-0.0223</td>
<td>1.94</td>
<td></td>
<td>1.87</td>
</tr>
<tr>
<td>0.05</td>
<td>0.0860</td>
<td>-0.2933</td>
<td>-0.1605</td>
<td>1.87</td>
<td></td>
<td>1.84</td>
</tr>
<tr>
<td>0.10</td>
<td>0.1735</td>
<td>-0.5933</td>
<td>-0.3323</td>
<td>1.84</td>
<td></td>
<td>1.83</td>
</tr>
<tr>
<td>0.20</td>
<td>0.3491</td>
<td>-1.1866</td>
<td>-0.6418</td>
<td>1.83</td>
<td></td>
<td>1.80</td>
</tr>
<tr>
<td>0.50</td>
<td>0.8045</td>
<td>-0.0277</td>
<td>-0.0017</td>
<td>2.80</td>
<td></td>
<td>2.51</td>
</tr>
<tr>
<td>1.00</td>
<td>0.0711</td>
<td>-0.1300</td>
<td>-0.0417</td>
<td>2.42</td>
<td></td>
<td>2.23</td>
</tr>
<tr>
<td>2.00</td>
<td>0.1514</td>
<td>-0.4596</td>
<td>-0.1839</td>
<td>2.24</td>
<td></td>
<td>2.20</td>
</tr>
</tbody>
</table>

Note that the van’t Hoff factor is more predictable for dilute solutions, as there is more recombination of the ions at higher concentrations.

Question:

Calculate the freezing point of 525g of water that contains 25.g of NaCl. Assume $i$, the van’t Hoff factor, is 1.85 for NaCl. ($K_{fp} = -1.86 \degree C/m$)

Answer:

$\Delta T = iK_{fp}m$

$\Delta T = (1.85)(-1.86)([25g/58.443g/mol]/.525kg) = -2.8 \degree C$

Osmosis and Osmotic Pressure

Osmosis: The movement of solvent molecules through a semipermeable membrane from a region of lower solute concentration to a region of higher solute concentration.
Two illustrations of osmosis

The height of the solution column provides a counter pressure, $\Pi$, which opposes the osmotic flow of the solvent.

Osmotic Pressure Formula

$\Pi = nRT/V$  
$\Pi = nRT/V$  
$\Pi = nRT/V$  

$n$ is the number of moles of solute and $\Pi$ is the osmotic pressure in atm.

$R$ is the ideal gas constant ($0.0821 \text{Latm/molK}$) and $V$ is the volume in L.

$\Pi = MRT$  

where $M$ is the molarity of the solution (mol/V)  
(the textbook gives it as $\Pi = cRT$)

For a 20% sucrose solution, the osmotic pressure is about 15atm, pushing a solution up 150m.

Determining Molar Mass of Solutes

For high mass molecules of solute (macromolecules) where the molarity can be low, osmotic pressure is the preferred method for determining molar mass.

For lower mass molecules (solute), freezing point depression (or BPE) is the preferred method.

Saline solutions (IV use)

0.92% NaCl (mass/vol) in blood cells

Solutions = 0.92% isotonic

>0.92% hypertonic (cell shrinks)

<0.92% hypotonic (cell bursts)

Reverse osmosis: Providing a pressure greater than $\Pi$ to reverse the flow of solvent particles can be used in water purification processes.

Colloids

Solutions: Relatively small particles; homogeneous mixture; no settling

Suspensions: Larger particles that can “settle out” (e.g. fine sand in water)

Colloids: High molar masses; relatively large diameters (~1,000nm); they do not “settle out”

<table>
<thead>
<tr>
<th>Type</th>
<th>Dispersing Medium</th>
<th>Dispersed Phase</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aerosol</td>
<td>Gas</td>
<td>Liquid</td>
<td>Fog, smoke, aerosol sprays</td>
</tr>
<tr>
<td>Aerosol</td>
<td>Gas</td>
<td>Solid</td>
<td>Smoke, airborne viruses, automobile exhaust</td>
</tr>
<tr>
<td>Foam</td>
<td>Liquid</td>
<td>Gas</td>
<td>Shaving cream, whipped cream</td>
</tr>
<tr>
<td>Foam</td>
<td>Solid</td>
<td>Gas</td>
<td>Shampoo, mousse</td>
</tr>
<tr>
<td>Emulsion</td>
<td>Liquid</td>
<td>Liquid</td>
<td>Mayonnaise, milk, face cream</td>
</tr>
<tr>
<td>Gel</td>
<td>Solid</td>
<td>Liquid</td>
<td>Jello, jell, chewee, butter</td>
</tr>
<tr>
<td>Gel</td>
<td>Liquid</td>
<td>Solid</td>
<td>Gold in water, rock of magnesium, mud</td>
</tr>
<tr>
<td>Solid sol</td>
<td>Solid</td>
<td>Solid</td>
<td>Milk glass</td>
</tr>
</tbody>
</table>
Tyndall Effect: Scattering of light from colloidal dispersions

Hydrophobic (water-fearing) Colloids
- Weak attractions
- Common with metals and nearly insoluble salts
- Precipitation of AgCl is usually a colloidal dispersion
- Generally do not coagulate into larger particles

Hydrophilic (water-loving) Colloids
- Often have –OH (hydroxyl) and –NH₂ (amine) groups
- Proteins, starch and homogenized milk are examples

Emulsions: Colloidal dispersions of one liquid in another
- e.g. oil in water
- Emulsions can be kept from separating by using an emulsifying agent (such as lecithin, a lipid found in egg yolks) that “links” the particles together
- Soaps and detergents are emulsifying agents: They have a hydrophobic and a hydrophilic end that links the nonpolar dirt and oil to polar water

Summary Questions
A physician studying a type of hemoglobin formed during a fatal disease dissolves 21.5mg of the protein in water at 5.0°C to make 1.50mL of solution in order to measure its osmotic pressure. At equilibrium, the solution has an osmotic pressure of 3.61torr. What is the molar mass of the hemoglobin?

Answer:
\[ Π = \frac{MRT}{V} \]
\[ M = \frac{ΠRT}{V} \]
\[ = \frac{(3.61\text{torr})}{(760\text{torr/atm})} \times \frac{(0.0821\text{Latm/molK})(273.15K + 5.0)}{0.00150\text{Lsoln}} = 2.08 \times 10^{-4} \text{M} \]
Moles solute = \( MV = (2.08 \times 10^{-4}\text{mol/Lsoln})(0.00150\text{Lsoln}) = 3.12 \times 10^{-7} \text{mol} \)
\[ 0.0215\text{g} / 3.12 \times 10^{-7} \text{mol} = 6.89 \times 10^{-4} \text{g/mol} \]

Using Raoult’s Law, explain what is happening in this series of before and after pictures.

Note that in figure (a), the beaker on the left is pure water, and the beaker on the right is a solution containing a nonvolatile solute.

Answer: The pure water will have a higher vapor pressure than the solution. The solution will try to reduce the vapor pressure by liquefying the excess water vapor, which will eventually “dry out” the pure water beaker.
Since a more volatile substance always has a greater vapor pressure relative to its number of moles (by definition), its mole fraction in the vapor state above the liquid is in a greater proportion relative to the mole fraction in solution where vapor pressure has no bearing.

Question: Above which of these solutions does the vapor pressure have the greater mole fraction of benzene at 25°C: A solution with equal masses of benzene \((\text{C}_6\text{H}_6)\) and toluene \((\text{C}_7\text{H}_8)\), or a solution with equal numbers of moles of benzene and toluene?

Benzene: 50% in solution and 77% in vapor
Toluene: 50% in solution and 23% in vapor

Answer: Because the molar mass of toluene is greater than that of benzene, equal masses will mean fewer moles of toluene in solution and a greater mole fraction of benzene. This means that the greater mole fraction of the vapor pressure will be for equal masses (the benzene is more volatile to begin with).

Question: A solution was prepared by dissolving 23.7g of \(\text{CaCl}_2\) in 375g of water. The density of the resulting solution was 1.05g/mL. What is the molarity of the \(\text{Cl}^-\) ion in this solution? Assume 100% dissociation of the compound. The molar mass of \(\text{CaCl}_2\) is 110.984g/mol.

Answer:
\[
\text{mol } \text{CaCl}_2 = \frac{23.7g}{110.984g/mol} = .21354 \text{mol } \text{CaCl}_2
\]
\[
\text{Mol } \text{Cl}^- = 2(\text{.21354}) = .42709 \text{mol } \text{Cl}^-
\]
\[
\text{Volume} = (23.7g + 375g) / (1 \text{mL} / 1.05g) = 3797L
\]
\[
.42709 \text{mol } \text{Cl}^- / .3797L = 1.12M
\]

Question: A 1.34m aqueous solution of compound X had a boiling point of 101.4°C. Which one of the following could compound X be: \(\text{CH}_3\text{CH}_2\text{OH}\), \(\text{C}_6\text{H}_12\text{O}_6\), \(\text{Na}_3\text{PO}_4\), \(\text{KCl}\), \(\text{CaCl}_2\)? The boiling point elevation constant for water is 0.52°C/m.

Answer:
\[
\Delta T_b = iK_b m
\]
\[
\Delta T_b = (0.52°C/m)(1.34m)
\]
\[
\Delta T_b = .697°C
\]
Since the actual \(\Delta T_b\) is 1.4°C (double the expected), \(\text{KCl}\) would be the only one close to the right answer.

Question: A 0.15m aqueous solution of a weak acid has a freezing point of -0.31°C. What is the percent ionization of this weak acid at this concentration? \(K_{fp}\) for water is -1.86°C/m.

Answer:
\[
-0.31°C = (-1.86°C/m)m
\]
\[
m = .167m
\]
\[
0.167/0.15 = 1.11 \text{ or } 11\% \text{ ionized}
\]