Chapter 18: Solutions
Part 1 - Notes: Homogeneous and Heterogeneous Aqueous Systems

Objectives: Explain the significance of the statement "Like dissolves like."
Distinguish between strong electrolytes, weak electrolytes, and non-electrolytes- giving examples of each.
Differentiate between colloids, suspension, and solutions on both macroscopic and microscopic levels.
Describe and explain the Tyndall effect and its significance.
Identify, define, and explain: aqueous solution, solvent, solute, solvation, electrolytes, non-electrolytes, strong electrolytes, weak electrolytes, suspension, colloid, Tyndall effect, Brownian motion, and emulsions.

Text Reference: Section 17.3 – pages 482-488 and Section 17.4 – pages 489-493

Solution: homogeneous and **stable** mixtures

Chemically pure water almost never exists in nature because . . .

Aqueous solutions:

Solvent:
For an aqueous solution, water is the solvent.

Solute:
Solutes may be ionic or molecular – and they are too small to be filtered out of a solution.

Sometimes, you put so much solute into water that some of the solute falls to the bottom of the container and sits there.
Equilibrium of a saturated solution with excess solute

Solvation:

Dissociation:

Solvation causes dissociation.

Dissociation of sodium bicarbonate:

Sometimes, the attractions between the ions in the crystal are stronger than the attractions between the ions and water. These compounds are not dissolved to any significant amount – their ions are not solvated. We say these are insoluble, like $CaCO_3$.

What types of compounds are generally soluble in water?

General rule: Polar solvents dissolve polar solutes and nonpolar solvents dissolve nonpolar solutes.

LIKE DISSOLVES LIKE!

Solubility Rules:

Miscible:

Immiscible:

Partially miscible:

Electrolyte:

Strong electrolyte:

Weak electrolyte:

Nonelectrolyte: compounds that do not conduct electric current in aqueous or molten states

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A solution is a homogeneous mixture – its components are evenly dispersed and a solution is stable. Let's no explore some heterogeneous aqueous systems: suspensions and colloids.

Suspension:

Colloids:

Tyndall effect:

Brownian Motion:

Let's make some comparisons:

Property	Solution	Colloid	Suspension
Particle type			
Particle size (approx)			
Effect of light			
Effect of gravity			
Filtration			
uniformity			

Gas dispersed in a liquid: **example:**

Gas dispersed in a solid: **example:**

Liquid dispersed in a liquid: **example:**

Liquid dispersed in a gas: **example:**

Solid dispersed in a gas: **example:**

Solid dispersed in a liquid: **example:**

Chapter 18: Solutions
Part 1 – Assignment: Homogeneous and Heterogeneous Aqueous Systems

1. Why does paint need to be stirred before it is used?

2. What is the basis for distinguishing between among solutions, suspensions, and colloids?

3. What is the Tyndall effect and how is Brownian motion related to the Tyndall effect?

4. Why is it recommended that drivers use low beam headlight when driving under foggy conditions at night?

5. Why is water an excellent solvent for ionic and polar compounds but not for nonpolar compounds?

6. Which of the following substances dissolve appreciably in water? Give reasons for your choice.
   a. HCl b. NaI c. NH_3 d. $MgSO_4$
   e. CH_4 f. $CaCO_3$ g. gasoline

7. What makes a colloidal dispersion stable?

8. Solutions do not demonstrate the Tyndall effect. Why not?

9. Explain why ethanol will dissolve in both gasoline (nonpolar) and water (polar).

10. For each description list all to which it applies: solution, suspension, or colloid. It may apply to more than one, list all.

    a. ____________________ does not settle out on standing
    b. ____________________ heterogeneous mixture
    c. ____________________ particle size less than 0.1 nm
    d. ____________________ demonstrates Tyndall effect
    e. ____________________ particles are invisible to the unaided eye
    f. ____________________ homogenized milk
    g. ____________________ salt water
    h. ____________________ jelly
    i. ____________________ mayonnaise

11. What is an emulsion and how are emulsions stabilized? (See text book section 17.4)

Chapter 18: Solutions
Part 2 - Notes: Properties of Solutions

Objectives: Identify the factors that determine the rate at which a solute dissolves.
Calculate the solubility of a gas in a liquid under various pressure conditions.
Identify, define, and explain: saturated solution, solubility, unsaturated, supersaturated, miscible, immiscible, and Henry's law.
Text Reference: Section 18.1 – pages 501-508

Mixture:

Mixtures may be ***heterogeneous*** or ***homogeneous***.
Heterogeneous:

Homogeneous:

A mixture is a ***homogeneous mixture.*** A mixture has at least two components: a solvent and at least one solute
The ***solute*** is the smaller component and the ***solvent*** is the larger component. You probably think of solutions as liquids – but that is not always the case. Let's think about different types of solutions . . .

	Solid – solute	***Liquid – solute***	***Gaseous - solute***
Solid – solvent			
Liquid – solvent			
Gas – solvent			

Alloy:

Let's make a solution. You have a sugar cube and you want to dissolve it in water. If you allow the sugar cube to sit in the water, it will dissolve. However, you are impatient and you want the sugar to dissolve in the water more quickly. How do you do this?

What factors make a solute dissolve more quickly in a solvent?

Why do these factors increase the rate at which the solute dissolves in the solvent?

Solubility:

You have 100 g of water at a temperature of 25^o and 1 atm of pressure. You dissolve 36.2 g of sodium chloride into the water. You stir and all the salt dissolves. The solution is clear.
You add 1 more grain of salt into the water – but it is unable to dissolve. What happens to it?

Saturated solution:

Unsaturated solution:

Supersaturated solution:

How do you make a supersaturated solution:

Testing for saturation levels:
Let's say you have a solution and the solute – and you want to test what level of saturation the solution has. What do you do?

Influence of temperature on solubility:
Temperature influences the rate at which a solute dissolves – but also the quantity of solute able to be dissolved.

Solid solute:

Gaseous solute:

Think of an unopened bottle of soda versus an opened bottle of soda.

Pressure is also a key factor of solvation when a gas is involved.

What is the result of an increase in pressure?

Henry's Law: at a given temperature, the solubility (S) of a gas in a liquid is directly proportional to the pressure (P) of the gas above the liquid.

$$S_1 / P_1 = S_2 / P_2$$

Example: The solubility of a gas in water is 0.77 g/L at 3.5 atm. What is its solubility at standard pressure, if the temperature is held at a constant 25°C?

Question: How is the solubility of a gas affected by pressure above the liquid?

Question: What could you do to change:

(a) a saturated solution to an unsaturated solution?

(b) an unsaturated solution to a saturated solution?

Question: Can a solution with undissolved solute on the bottom be supersaturated?

Chapter 18: Solutions
Part 2 – Assignment: Properties of Solutions

1. How does the solution process occur and what influences it?

2. List the three factors that affect the solubility of a solid substance. Of those three factors, which affect the solubility of a substance?

3. Why does a dissolved component not settle out of solution?

4. If a solution of saturated sodium nitrate is cooled, what change might you observe?

5. The solubility of methane, the major component of natural gas, in water at 20°C and standard pressure is 0.026 g/L. If the temperature remains constant, what is the solubility of the gas at (a) 0.60 atm and (b) 1.80 atm?

6. You are given a clear aqueous solution containing KNO_3. How would you determine experimentally if the solution is saturated, unsaturated, or supersaturated?

7. A solution contains 26.5 g NaCl in 75.0 g of water at 20°C. Determine if the solution is saturated, unsaturated, or supersaturated if the solubility for NaCl at 20° is 36.0 g/100 g H_2O.

8. How many grams of $NaNO_3$ will precipitate if a saturated solution of $NaNO_3$ in 200.0 g of water is saturated at 50°C and then cooled to 20°C? $NaNO_3$ solubility: @ 50°C = 114.0 g/100g H_2O @ 20°C = 88.0g/100g H_2O

9. Why does increasing the temperature allow for less gaseous solute to be dissolved?

10. Why does increasing the temperature allow for more solid solute to be dissolved?

Chapter 18: Solutions
Part 3 – Notes: Concentrations of Solutions

Objectives: Solve problems involving molarity, molality, percent by mass, percent by volume, parts per hundred, parts per million, and parts per billion of a solution.
Describe how to prepare dilute solutions from concentrated solutions of known concentration.
Explain what is meant by percent by volume (% (v/v)) and percent by mass (%(m/v)) solutions.
Identify, define, and explain: molarity, molality, parts per hundred, parts per million, parts per billion, and mole fraction.

Text Reference: Section 18.2 – pages 509-515 and 18.4 (part) – pages 520-522

Qualitative versus Quantitative: Concentrated versus dilute = qualitative
Science needs a quantitative statement of concentration.

Solubility may be measured in terms of parts per million (ppm), parts per billion (ppb), grams of solute per 100 g of solvent, MOLARITY, and MOLALITY.

Molarity: the number of ***moles*** of solute per ***liter of solution***. It is important to note that the solution is equal to the solute plus the solvent. Unless otherwise noted, the solvent used for our work will be water.

MOLARITY =

Example 1: *What is the molarity of a solution made with 126.32 g of sodium hydroxide dissolved to make 874.2 mL of a solution? (Remember to show all the required work, units, etc.)*

Molality: the number of ***moles*** of solute per ***kg of solvent***. It is important to remember that the solution is equal to the mass of the mass of the solute plus the mass of the solvent. In molality, the mass of interest is that of the solvent alone.

MOLALITY =

Example 2: What is the concentration, in molality, of a solution made with 145.8 g ammonium carbonate dissolved in 845.6 g of water?

Example 3: What is the mass, in grams, of copper (II) chloride present in 741.2 g of water to make a 2.56 m solution?

Example 4: What is the mass of the ***solution*** made from 125.0 g of sodium nitrate with enough solvent to make a 1.75 m solution? Recall: solution = solute + solvent.

Percent, Parts per million and Parts per billion

Remember, percent is also called parts per hundred. These problems are very similar.
This is a per mass concentration. This may also be done with a per volume concentration.
These calculations may be mass/mass (m/m), mass/volume (m/v), or volume/volume (v/v).

Percent = pph =

ppm =

ppb =

Example 5: What is the concentration, in ppm of a solution made with 18.5 g of salt in 12 500.0 g of water?

Example 6: How many grams of glucose ($C_6H_{12}O_6$) would you need to prepare 2.0L of 2.8% glucose (m/v) solution?

Dilution

Sometimes you have a concentrated solution or one with a high concentration (*M* or *m*) and you want to make a more dilute solution. This is how acid solutions used in lab are made. The chemical stock room has reagent grade high concentrated acids in bottles. To use them in a high school lab, they need to be diluted – made less concentrated.

What does this mean to take a concentrated solution and dilute in terms of solute, solvent, and solution?

How do you go about this? Let's say you want to make 250.0 mL of a 1.5*M* HCl solution. All you have in the stock room is 12.0*M* HCl. How much of the more concentrated solution do you need to mix with additional water to make the dilute solution?

Find the number of moles of solute that you want to have in your final solution:

Now find the volume of concentrated acid you would need to get that quantity of solute:

This could have been done in a slightly more simple manner. Since you want the number of moles to be the same in both solutions, you can rearrange the molarity equation to solve for moles and then set that equal to itself.

$$M_1V_1 = M_2V_2$$

Example: How many milliliters of a stock solution of 2.00*M* $MgSO_4$ would you need to make 100.0 mL of 0.400*M* $MgSO_4$?

Chapter 18: Solutions
Part 3 - Assignment: Concentrations of Solutions

Solve the following problems. Be sure to show all required work, units, formulas, etc.

1. If 162.35 g aluminum hydroxide are dissolved in 6750 mL of solution, what is the molarity of the solution?

2. Calculate the molality of a solution containing 0.0762 mol of molecular iodine in 450.0 g of carbon tetrachloride.

3. You have 125 g of potassium sulfate and 385 g water. You mix them together to make a solution.
   a. What is the molality of the solution you make?

   b. What is the molarity of the solution you make? (Recall: solution = solute + solvent. For this: 1 g = 1 mL.)

4. A bottle of hydrogen peroxide antiseptic is labeled 3.0% (v/v). How many mL H_2O_2 are in a 400.0-mL bottle of this?

5. What is the concentration, in percent (m/v), of a solution with 75 g potassium sulfate in 1500 mL of solution?

6. Hydrogen peroxide is often sold commercially as a 3.0% (m/v) aqueous solution.
   a. If you buy a 250.-mL bottle of 3.0% hydrogen peroxide (m/v), how many grams of H_2O_2 have you purchased?

   b. What is the molarity of this solution?

7. How many milliliters of 1.50*M* nitric acid contain enough acid to dissolve a 3.94g old copper penny?

$$3Cu + 8\ NHO_3 \rightarrow 3Cu(NO_3)_2 + 2\ NO + 4H_2O$$

8. How much stock H_2SO_4 (18*M*) would you need to make 375 mL of a 2.35*M* solution?

Chapter 18: Solutions
Part 4 – Notes: Colligative Properties and Calculations

Objectives: Explain, on a microscopic level, the change in vapor pressure of a solution when compared to the pure solvent.
Explain, on a microscopic level, the elevation of the boiling point of a solution and the depression of the freezing point of a solution when compared to the pure solvent.
Calculate the molar mass of a molecular compound from the freezing-point depression or boiling point elevation.
Caclulate the boiling-point elevation or freezing-point depression of a solution – from the base solvent.
Identify, define, and explain: colligative property, boiling-point elevation, freezing-point depression, molal boiling-point elevation constant, and molal freezing-point depression constant.

Text Reference: Section 18.3 – pages 517-519 and Section 18.4 (Part) – pages 522-522

Colligative property:

Key colligative properties are . . .

Vapor pressure is the pressure exerted by a vapor that is in dynamic equilibrium with its liquid in a closed system.

A solution that contains a solute that is nonvolatile (not easily vaporized) always has a lower vapor pressure than the pure solvent.

Why:

The addition of a solute causes there to be a decrease in vapor pressure. If a substance has a lower vapor pressure, then it must be raised to a greater temperature to get the solution to boil. So . . .

A solution has a boiling point that is higher than the pure solvent.

Why:

Boiling-point elevation:

The addition of a solute also causes a change in freezing point.

A solution has a lower freezing point than the pure solvent.

Why:

Freezing-point depression:

Recall: **Mole Fraction:**

The size and type of the molecules or ions that compose the solute **do NOT** determine how it will affect the boiling point or freezing point of a solution, but rather it is the **number of dissolved ions or molecules** that affect the boiling point and freezing point of a solution.

Example 1: What is the concentration of the sucrose particles in water, in molality, if 1 mole of sucrose ($C_{12}H_{22}O_{11}$) is dissolved in 1 kg of water?

Molality of sucrose: *Molality of dissolved particles:*

Note: Sucrose is a non-electrolyte – it does not dissociate in water.

Example 2: What is the concentration of the sodium chloride particles in water, in molality, if 1 mole of sodium chloride is dissolved in 1 kg of water?

Molality of sodium chloride: *Molality of dissolved particles:*

Note: NaCl is an ionic compound and it easily dissociates into **TWO** ions in aqueous solution.

Example 3: What is the concentration of the calcium chloride particles in water, in molality, if 1 mole of calcium chloride is dissolved in 1 kg of water?

Molality of calcium chloride: *Molality of dissolved particles:*

Note: $CaCl_2$ is an ionic compound and it easily dissociates into **THREE** ions in aqueous solution.

There are three times as many dissolved particles hydrated when calcium chloride is dissolved than when sucrose is dissolved. Therefore, the calcium chloride affects boiling and freezing points of the solution three times more than does the sucrose.

Calculations Involving Boiling Point and Freezing Point Changes

Boiling Point Elevation

ΔT_b if the difference between the boiling point of the pure solvent and that of the solution. For our purposes, there is always an increase in the boiling point; it is elevated.

$\Delta T_b =$ $\Delta T_b =$

k_b is the constant that relates change in T to the molality of the dissolved particles. k_b is the ***molal boiling point constant***.
Note: the k_b constant is different for every solvent. $k_{b\,(water)} =$

$\mathbf{m_{dp}}$ is the molality of the dissolved particles.
Remember: The molality of the ionic compound is different from the molality of dissolved particles.

$\mathbf{m_{dp}} =$

Freezing Point Depression

ΔT_f is the difference between the freezing point of the pure solvent and that of the solution. For our purposes, there is always a decrease in the freezing point; it is depressed.

$\Delta T_f =$ $\Delta T_f =$

k_f is the constant that relates change in T to the molality of the dissolved particles. k_f is the ***molal freezing point constant***.
Note: the k_f constant is different for every solvent. $k_{f\,(water)} =$

Example 4: Calculate the boiling point for each of the solutions in examples 1and 3 above.

Example 5: What is the freezing point of a 2.35*m* solution of sodium chloride?

Chapter 18: Solutions
Part 4 – Assignment: Colligative Properties and Calculations

Use the information in the data table to answer questions about BP and FP changes.

Solvent	*Normal FP (°C)*	*k_f (°C/m)*	*Normal BP (°C)*	*k_b (°C/m)*	*Formula*
benzene	5.50	5.1	80.15	2.53	C_6H_6
naphthalene	80.20	6.9	218.0	5.65	$C_{10}H_8$
phenol	40.90	7.1	181.2	3.56	C_6H_5OH
water	0.00	1.86	100.0	0.52	H_2O

Note: Sucrose = $C_{12}H_{22}O_{11}$ = 342.34 g/mol   Glucose = $C_6H_{12}O_6$ = 180.18 g/mol

1. A solution is prepared by dissolving 6.75 g glucose in 325 g water. What is the freezing point of the solution?

2. A 0.400 m solution of naphthalene in benzene is needed. 32.0 g naphthalene is available. How much benzene solvent is needed? What is the boiling point of this solution? (Naphthalene does not dissociate.)

3. 65.43 g ammonium nitrate are dissolved in 654.3 g water. What is the freezing point of the solution?

4. A 5.250 g sample of a newly synthesized, non-dissociating compound is added to 250.0 g water and the freezing point of the solution was 0.62°C. What is the molar mass of the new compound?

5. A solution contains 98.76 g sodium carbonate dissolved in 765.4 g water. What is the boiling point of the solution?