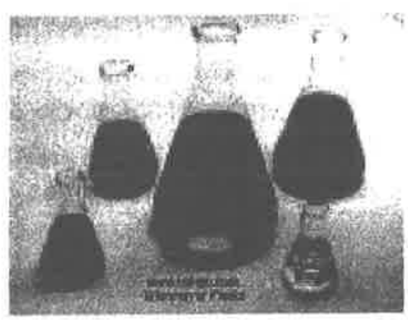
Solutions



Honors Chemistry

Molality Problems

Molality = moles of solute/kilogram of solvent
M =m/kg
If both substances are the same state, the one in the greater quantity is the solute.

1. What is the molality of a solution made with 80.0 grams of NaOH and 855 grams of water?

2. Calculate the molality of a copper alloy made from 35.0 g of silver and 100.0 g of copper.

3. Calculate the number of grams of solute necessary to prepare 700.0g of an aqueous 0.500m solution of sulfuric acid.

4. Calculate the number of grams of C_2H_5OH placed in 750.0 grams of water to create a 2.00molal solution.

5. Calculate the number of grams of water that must be added to 65.0 grams of glucose, $C_6H_{12}O_6$, in the preparation of a 3.00m solution.

6. Calculate the number of grams of water that must be added to 4.10 mol of sulfuric acid in the preparation of a 12.0m solution.

Dilutions

Solutions of known molarity are often available in the laboratory. You can use these to make new solutions of lower concentrations by just adding more solvent. As you observed in the inquiry molarity lab activity, *the number of moles of solute does not change when a solution is diluted.*

MOLARITY (M) = **moles of solute** divided by **liters of solution (solute + solvent)**
If we rearrange the equation...then...

moles solute = Molarity x Liters of solution

In a dilution, the ***moles of solute before dilution = moles solute after dilution***
Substitute from the previous equation...

Molarity x Volume before dilution = Molarity x Volume after dilution

$$M_1V_1 = M_2V_2$$

M_1 and V_1 represent the initial solution's molarity and volume (before dilution), M_2 and V_2 represent the solution's final molarity and volume after dilution. Volume can be in either milliliters or liters as long as the units are the same for both V_1 and V_2.

You also need to be able to explain how to prepare a diluted solution in a sentence or two if asked. See example 1 below.

Example 1:
A student needs to prepare 100.0 ml of 0.50M NaOH. He has a stock of 3.0M NaOH. How does he prepare the desired solution?
First determine the volume of 3.0M NaOH needed: (3.0M)(V1)? = (0.50M)(0.100L)

V1 = 0.017 L or 17 mL

"Measure 17 ml of 3.0M NaOH, transfer to a volumetric flask and add water until the final volume is 100.0 ml."

Example 2
A student adds water to 25 ml of 4.0 M NaOH until the final volume is 40.0 ml. What is the molarity of the new solution?

4.0M x 0.025L = ? x 0.040L
? = 2.5 M

Concept Practice:
A. Explain why, after a dilution is completed, the resulting solution is always of lower concentration than the original.

B. Your experiment requires 5 ml of 1.0M KOH. You have 1 L of 0.5M KOH on the shelf. Can you prepare the required solution? Explain your answer.

Problems:

1. How do you prepare 250.0 ml of 0.20 M NaCl from a stock of 1.0M NaCl?

2. How do you prepare 400.0 ml of 1.0M $Zn(NO_3)_2$ from a stock of 2.5M $Zn(NO_3)_2$?

3. What is the resulting concentration when 400.0 ml of 12 M H_2SO_4 is diluted to a new volume of 500.0 ml?

4. What is the resulting concentration when 20.0 ml of 6.0M HCl is added to 100.0 ml of water? (Tricky!!!)

5. What volume of 2.0 M LiCl is needed to make 500.0 ml of 0.35M LiCl?

6. What is the molarity of a solution made by adding 50.0 ml water to 150.0 ml of 0.10M KOH?

7. What volume of 12.0 M HCl is needed to prepare 450. ml of 2.0 M HCl? How much water is needed?

8. Determine the concentration of a solution made by diluting 75ml of 5.0M NaOH with 200.0 ml of water.

9. Determine the concentration of stock solution used when 40.0 ml of it are needed to make 800.0 ml of 0.20 M $HC_2H_3O_2$.

10. Determine the volume of 1.0M KBr needed to prepare 500.0 ml of 0.20 M KBr.

Solution Math

1. What is the molarity of a solution made by dissolving 23.0 grams NaCl in enough water to make 40.0 mL of solution?

2. What mass of KNO_3 is needed to make 2.0 L of a 0.20 molar solution?

3. Determine the percent by mass of solute in a solution containing 134g of $Pb(NO_3)_2$ in 266g of water? Hint: Total needs to account for solute + solvent mass.

4. How many grams of lead(II) nitrate are there in 500.0 mL of a 0.25 M solution?

5. What is the molarity of a solution made by dissolving 0.66 moles of NaI in enough water to make 5.0 liters of solution?

6. How many grams of water are needed to make a 25% by mass hydrochloric acid solution using 200g of HCl?

7. How many moles are there in 50.0 mL of a 0.002 M solution of NaCl?

8. Accurately explain how to prepare 250.0 mL of a 0.20 M silver nitrate solution.

Use the equation below to answer the following questions.

$$K_4FeC_6N_{6\,(aq)} + 2\ Zn(NO_3)_{2\,(aq)} \dashrightarrow 4\ KNO_{3\,(aq)} + Zn_2FeC_6N_{6\,(s)}$$

9. What volume of 2.5 M zinc nitrate is needed to produce 1.8g of zinc ferrocynide?

10. If 20.5 mL of 0.12 M potassium ferrocynide and 15.8 mL of 0.20 M zinc nitrate react, how many grams of zinc ferrocynide is formed?

11. Determine the volume of 0.3 M zinc nitrate needed to completely react with 34.9 mL of 0.50 M potassium ferrocynide?

Use the equation below to answer the following questions.

$$K_2CrO_{4\ (aq)} + 2\ AgNO_{3\ (aq)} \dashrightarrow 2\ KNO_{3\ (aq)} + Ag_2CrO_{4\ (s)}$$

12. Determine the number of moles of potassium nitrate produced when 67mL of 1.2 M potassium chromate react with 56 mL of 0.98 M silver nitrate.

13. Determine the volume of .070 M silver nitrate needed to precipitate out all the CrO_4^{-2} ions from 22 mL of 0.30 M potassium chromate.

14. Determine the mass of silver chromate produced from 84 mL of 1.3 M silver nitrate.

15. What volume of 0.25 M potassium chromate is needed to completely react with 15.6 g of silver nitrate?

Use the equation below to answer the following questions.

$$2\ Al(NO_3)_{3\ (aq)} + 3\ (NH_4)_2S_{\ (aq)} \dashrightarrow Al_2S_{3\ (s)} + 6\ NH_4NO_{3\ (aq)}$$

16. Determine the volume of 0.40 M ammonium sulfide needed to completely react with 9.6g of aluminum nitrate.

17. Determine the mass of aluminum sulfide produced from 112 mL of 3.4 M $(NH_4)_2S$.

18. How many moles of NH_4NO_3 are produced from 12.3 mL of 0.10 M $(NH_4)_2S$ reacting with 15.2 mL of 0.15 M $Al(NO_3)_3$?

Solution Math Review

1. Accurately explain how to prepare 2.0 L of a 1.5 M copper (II) nitrate solution.

2. How much water is needed to make 3.0 L of 35% HCl solution if the density of the resulting solution is 1.1g/ml?

3. How many grams of sodium hydroxide must be weighed out to make 2.0L of 3.0M NaOH solution?

4. What is the % by mass of a solution made by dissolving 35.8 grams of silver nitrate in 500. grams of water?

5. What is the molarity of 2.56 moles of potassium iodide dissolved in 250. mL of solution?

Use the equation below to answer the following questions.

$$Na_2CO_{3\,(aq)} + 2\ AgNO_{3\,(aq)} \dashrightarrow 2\ NaNO_{3\,(aq)} + Ag_2CO_{3\,(s)}$$

6. Determine the volume of 0.10 M silver nitrate needed to completely react with 125 mL of 0.15 M sodium carbonate?

7. Calculate the mass of Ag_2CO_3 formed when 45 mL of 1.2 M silver nitrate reacts completely?

8. What volume of 1.5 M silver nitrate is needed to produce 10.4g of Ag_2CO_3?

9. If 40.5 mL of 0.15M silver nitrate reacts with 30.0 mL of 0.12 M Na_2CO_3, what mass of Ag_2CO_3 is formed?

Use the equation below to answer questions 10-14.

$$2\ Al(NO_3)_3\ _{(aq)} + 3\ (NH_4)_2S\ _{(aq)} \dashrightarrow Al_2S_3\ _{(s)} + 6\ NH_4NO_3\ _{(aq)}$$

10. Determine the volume of 0.80 M ammonium sulfide needed to completely react with 150.0 mL of 0.50 M aluminum nitrate.

11. Determine the mass of Al_2S_3 produced when 450.0 mL of 3.5 M $(NH_4)_2S$ reacts completely.

12. How many moles of Al_2S_3 are produced from 125 mL of 0.30 M $Al(NO_3)_3$ mixed with 100.0 mL of 0.50 M $(NH_4)_2S$?

13. Determine the number of moles of NH_4NO_3 produced when 425mL of 3.0 M $Al(NO_3)_3$ reacts completely.

14. How many moles of aluminum nitrate are needed to completely react with 0.5 L of 0.25 M ammonium sulfide?

15. If you prepared a saturated solution of sodium nitrate at room temperature and then heated it to about 75°C, would the solution still be saturated? Explain.

16. Explain why carbonated beverages become flat when left opened.

Name ______________________________ Date______________

Solutions: Chapter 12

Solution: Homogeneous ____________ in a single phase.

Solvent: The substance that does the ______________.

Solute: The substance that gets ________________.

Soluble: Capable of being dissolved in a particular ___________. Example: Sodium chloride is soluble in water.

Insoluble: Incapable of being dissolved in a particular solvent. Example: Sodium chloride is insoluble in hexane. What is the special name for a solution in which the solute and the solvent are solids?

Miscible: Liquids that dissolve freely in one another in all ______________. Example: ____________________

Immiscible: Liquids that do not dissolve in one another. (Oil and water)

Solvation: The process of _________________

What is an example of a heterogeneous mixture?

Solute	Solvent	Types of solutions	Examples
Solid	Solid	Solid in solid	Alloys
Liquid	Solid	Liquid in solid	Hydrated salts
Gas	Solid	Gas in solid	Dissolved gases in minerals
Solid	Liquid	Solid in liquid	Salt solution in water
Liquid	Liquid	Liquid in liquid	Alcohol in water
Gas	Liquid	Gas in liquid	Aerated drinks
Solid	Gas	Solid in gas	Iodine vapours in air
Liquid	Gas	Liquid in gas	Humidity in air
Gas	Gas	Gas in gas	Air

What is the name of a substance the looks like a true solution , however, its' particles are not truly dissolved in the solvent?

If more energy is needed to separate than is released, the enthalpy is a positive value The solution would get colder.

If less energy is needed to separate compared to the release of energy, the solution will get warmer. The enthalpy value would be negative. (Exothermic)

Write the formula for molarity below:
How many grams of NaCl are required to make 0.500L of 0.25M NaCl?

Find the molarity of a 250 mL solution containing 10.0 g of NaF.

The amount of solute in a solution.
Describing Concentration
•% by mass - medicated creams
•% by volume - rubbing alcohol
•ppm, ppb - water contaminants
•molarity - used by chemists
•molality - used by chemists

Regulated Substances

Substance	Test Year	Concentration Range Found in SAWS Water	Highest Concentration Found in SAWS Water	MCL	MCLG	Possible Source
Nitrate (ppm)	1999	1.53 - 1.94	1.94	10	10	Runoff from fertilizer use; leaching from septic tanks, sewage; erosion of natural deposits
Barium (ppm)	1999	0.036 - 0.090	0.090	2	2	Discharge from drilling wastes; discharge from metal refineries; erosion of natural deposits.
Fluoride (ppm)	1999	0.1 - 0.3	0.3	4	4	Erosion of natural deposits; Discharge from fertilizer and aluminum factories
Antimony (ppb)	1999	2.7	2.7	6	6	Discharge from petroleum refineries; fire retardants; ceramics; electronics; solder
Tetrachloroethylene (ppb)	1999	0.5 - 1.2	1.2	5	0	Leaching by PVC pipes; discharge from factories and dry cleaners
Total Trihalomethanes (ppb)	1999	20.7	20.7	100	N/A	By-products of drinking water chlorination
Methylene Chloride (ppb)**	1999	0.5	0.5	5	0	Discharge from pharmaceutical and chemical factories

Write the formula to calculate molality in the space below:

Find the molality of a solution containing 75 g of $MgCl_2$ in 250 mL of water.

How many grams of NaCl are req'd to make a 1.54*m* solution using 0.500 kg of water?

Dilution problems:

$$\boxed{M_1V_1 = M_2V_2}$$

What volume of 15.8*M* HNO_3 is required to make 250 mL of a 6.0*M* solution?

Name ______________________________ Date__________

Molarity problems:

As is clear from its name, **molarity** involves moles.

The molarity of a solution is calculated by taking the moles of solute and dividing by the liters of solution.

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

This is probably easiest to explain with examples.

Example #1 - Suppose we had 1.00 mole of sucrose (342.3 grams) and proceeded to mix it into some water. It would dissolve and make sugar water. We keep adding water, dissolving and stirring until all the solid was gone. We then made sure that when everything was well-mixed, there was exactly 1.00 liter of solution.

What would be the molarity of this solution?

$$\text{Molarity} = \frac{1.00\text{ mol}}{1.00\text{ L}}$$

The answer is 1.00 mol/L. Notice that both the units of mol and L remain. Neither cancels.

And never forget this: replace the M with mol/L when you do calculations. The M is just shorthand for mol/L.

Example #2 - Suppose you had 2.00 moles of solute dissolved into 1.00 L of solution. What's the molarity?

$$\text{Molarity} = \frac{2.00\text{ mol}}{1.00\text{ L}}$$

The answer is 2.00 M.

Notice that no mention of a specific substance is mentioned at all. The molarity would be the same. It doesn't matter if it is sucrose, sodium chloride or any other substance. One mole of anything contains 6.022×10^{23} units.

Example #3 - What is the molarity when 0.75 mol is dissolved in 2.50 L of solution?

$$\text{Molarity} = \frac{0.75\text{ mol}}{2.50\text{ L}}$$

The answer is 0.300 M.

Now, let's change from using moles to grams. This is much more common. After all, chemists use balances to weigh things and balances give grams, NOT moles.

Example #4 - Suppose you had 58.44 grams of NaCl and you dissolved it in exactly 2.00 L of solution. What would be the molarity of the solution?

The solution to this problem involves two steps which will eventually be merged into one equation.

Step One: convert grams to moles.

Step Two: divide moles by liters to get molarity.

In the above problem, 58.44 grams/mol is the molar mass of NaCl.

Dividing 58.44 grams by 58.44 grams/mol gives 1.00 mol.

Then, dividing 1.00 mol by 2.00 L gives 0.500 mol/L (or 0.500 M).

Complete examples #5 , #6, #7, #8 (Ask your instructor for the answers when you are finished.)

5) Calculate the molarity of 25.0 grams of KBr dissolved in 750.0 mL.

6) 80.0 grams of glucose ($C_6H_{12}O_6$, mol. wt = 180. g/mol) is dissolved in enough water to make 1.00 L of solution. What is its molarity?

7) Calculate the molarity when 75.0 grams of $MgCl_2$ is dissolved in 500.0 mL of solution.

8) How many grams of $KMnO_4$ are needed to make 500.0 mL of a 0.200 M solution?

Molality

As is clear from its name, molality involves moles.

The molality of a solution is calculated by taking the moles of solute and dividing by the kilograms of solvent.

$$\boxed{\text{Molality} = \frac{\text{moles of solute}}{\text{kilograms of solvent}}}$$

This is probably easiest to explain with examples.

Example #1 - Suppose we had 1.00 mole of sucrose (it's about 342.3 grams) and proceeded to mix it into exactly 1.00 liter water. It would dissolve and make sugar water. We keep adding water, dissolving and stirring until all the solid was gone. We then made sure everything was well-mixed.

What would be the molality of this solution? Notice that my one liter of water weighs 1000 grams (density of water = 1.00 g / mL and 1000 mL of water in a liter). 1000 g is 1.00 kg, so:

$$\boxed{\text{Molality} = \frac{1.00 \text{ mol}}{1.00 \text{ kg}}}$$

The answer is 1.00 mol/kg. Notice that both the units of mol and kg remain. Neither cancels. A replacement for mol/kg is often used. It is a lower-case m and is often in italics, ***m***. However, if you write 1.00 m for the answer, without the italics, then that usually is correct because the context calls for a molality. Having said that, however, be aware that often m is used for mass, so be careful.

When you say it out loud, say this: "one point oh oh molal."

And never forget this: replace the m with mol/kg when you do calculations. The m is just shorthand for mol/kg.

Example #2 - Suppose you had 2.00 moles of solute dissolved into 1.00 L of solvent. What's the molality?

$$\boxed{\text{Molality} = \frac{2.00 \text{ mol}}{1.00 \text{ kg}}}$$

The answer is 2.00 m.

Notice that no mention of a specific substance is mentioned at all. The molarity would be the same. It doesn't matter if it is sucrose, sodium chloride or any other substance. One mole of anything contains 6.022×10^{23} units.

Example #3 - What is the molality when 0.75 mol is dissolved in 2.50 L of solvent?

$$\boxed{\text{Molality} = \frac{0.75 \text{ mol}}{2.50 \text{ kg}}}$$

The answer is 0.300 m.

Example #4 - Suppose you had 58.44 grams of NaCl and you dissolved it in exactly 2.00 kg of pure water (the solvent). What would be the molality of the solution?

The solution to this problem involves two steps.

Step One: convert grams to moles.

Step Two: divide moles by kg of solvent to get molarity.

In the above problem, 58.44 grams/mol is the molecular weight of NaCl. (For you technical types, I know it actually is a formula weight, but I'm glossing over the difference for the time being. Remember, this is a high school tutorial.)

Dividing 58.44 grams by 58.44 grams/mol gives 1.00 mol.

Then, dividing 1.00 mol by 2.00 kg gives 0.500 mol/kg = 0.500 m.

Complete the following problems. When you are finished, see your instructor for the answers.

#5) Calculate the molality of 25.0 grams of KBr dissolved in 750.0 mL water.

#6) 80.0 grams of glucose ($C_6H_{12}O_6$, mol. wt = 180. g/mol) is dissolved in 1.00 kg of solvent. What is its molalty?

#7) 100.0 grams of sucrose ($C_{12}H_{22}O_{11}$, mol. wt. = 342.3 g/mol) is dissolved in 1.50 L of water. What is the molality?

How to Calculate Concentration

: **concentration** of a chemical solution refers to the amount of solute that is dissolved in a solvent. Although . common to think of a solute as a solid that is added to a solvent (e.g., adding table salt to water), the solute could exist in another phase. If the solute and the solvent are in the same phase, then the solvent is the substance presence in the largest percentage. For example, if we add a small amount of ethanol to water, then the ethanol is the solute and the water is the solvent. If we add a smaller amount of water to a larger amount of ethanol, then the water would be the solute.

Units of Concentration

Once the solute and solvent have been identified, you can determine the concentration of the solution. There are several ways to express concentration. The most common units are **percent composition by mass, mole fraction, molarity, molality,** or **normality**.

1. **Percent Composition by Mass (%)**

   This is the mass of the solute divided by the mass of the solution (mass of solute plus mass of solvent), multiplied by 100.

   Example:
   Determine the percent composition by mass of a 100 g salt solution which contains 20 g salt.

   Solution:
   20 g NaCl / 100 g solution x 100 = 20% NaCl solution

2. **Mole Fraction (X)**

   This is the number of moles of a compound divided by the total number of moles of all chemical species in the solution. The sum of all mole fractions in a solution must equal 1.

   Example:
   What are the mole fractions of the components of the solution formed when 92 g glycerol is mixed with 90 g water? (molecular weight water = 18; molecular weight of glycerol = 92)

   Solution:
   90 g water = 90 g x 1 mol / 18 g = 5 mol water
   92 g glycerol = 92 g x 1 mol / 92 g = 1 mol glycerol
   total mol = 5 + 1 = 6 mol
   x_{water} = 5 mol / 6 mol = 0.833
   $x_{glycerol}$ = 1 mol / 6 mol = 0.167
   It's a good idea to check your math by making sure the mole fractions add up to 1:
   $x_{water} + x_{glycerol}$ = .833 + 0.167 = 1.000

3. **Molarity (M)**

   Molarity is probably the most commonly used unit of concentration. It is the number of moles of solute per liter of solution (not necessarily the same as the volume of solvent!).

   Example:
   What is the molarity of a solution made when water is added to 11 g $CaCl_2$ to make 100 mL of solution?

Solution:
11 g $CaCl_2$ / (110 g $CaCl_2$ / mol $CaCl_2$) = 0.10 mol $CaCl_2$
100 mL x 1 L / 1000 mL = 0.10 L
molarity = 0.10 mol / 0.10 L
molarity = 1.0 M

4. **Molality (m)**

Molality is the number of moles of solute per kilogram of solvent. Because the density of water at 25°C is about 1 kilogram per liter, molality is approximately equal to molarity for dilute aqueous solutions at this temperature. This is a useful approximation, but remember that it is only an approximation and doesn't apply when the solution is at a different temperature, isn't dilute, or uses a solvent other than water.

Example:
What is the molality of a solution of 10 g NaOH in 500 g water?

Solution:
10 g NaOH / (4 g NaOH / 1 mol NaOH) = 0.25 mol NaOH
500 g water x 1 kg / 1000 g = 0.50 kg water
molality = 0.25 mol / 0.50 kg molality = 0.05 M / kg molality = 0.50 m

5. **Normality (N)**

Normality is equal to the *gram equivalent weight* of a solute per liter of solution. A gram equivalent weight or equivalent is a measure of the reactive capcity of a given molecule. Normality is the only concentration unit that is reaction dependent.

Example:
1 M sulfuric acid (H_2SO_4) is 2 N for acid-base reactions because each mole of surfuric acid provides 2 moles of H^+ ions. On the other hand, 1 M sulfuric acid is 1 N for sulfate precipitation, since 1 mole of sulfuric acid provides 1 mole of sulfate ions.

Making Dilutions

You **dilute** a solution whenever you add solvent to a solution. Adding solvent results in a solution of lower concentration. You can calculate the concentration of a solution following a dilution by applying this equation:

$M_iV_i = M_fV_f$

where M is molarity, V is volume, and the subscripts i and f refer to the initial and final values.

Example:
How many millilieters of 5.5 M NaOH are needed to prepare 300 mL of 1.2 M NaOH?

Solution:
$5.5\ M \times V_1 = 1.2\ M \times 0.3\ L$ $V_1 = 1.2\ M \times 0.3\ L / 5.5\ M$
$V_1 = 0.065\ L$ $V_1 = 65\ mL$

So, to prepare the 1.2 M NaOH solution, you pour 65 mL of 5.5 M NaOH into your container and add water to get 300 mL final volume.

olutions are Mixtures

1. ***Solutions*** are stable, homogeneous mixtures
   a. Figure 11-1: adding ocean salt crystals to fish tank – particles of each substance are evenly dispersed.
2. Suspension particles will settle
   a. ***Suspension*** – mixture that appears uniform while being stirred, but separates into different phases when agitation ceases
      i. Example: Figure 11-2: muddy water – NOT a solution ingredients show visible cloudiness and separate spontaneously clay does not dissolve in water

Describing Solutions

1. Solute dissolves in the solvent
   a. ***Solvent*** – material dissolving the solute to make the solution
   b. ***Solute*** – material dissolved in a solution
2. Most solutions have a solid solute and liquid solvent … BUT
   a. ***Miscible*** – liquids or gases that will dissolve in each other
      i. Example: dissolved CO_2 in a soda bottle
   b. ***Immiscible*** – liquids or gasses that will NOT dissolve in each other
      i. Example: Figure 11-3: lava lamp has 2 immiscible liquids
3. Solubility is the maximum that can dissolve
   a. ***Soluble*** – can be dissolved in a particular solvent
   b. ***Insoluble*** – does not dissolve much in a particular solvent
      i. Example: Figure 11-5: Temperature/Solubility Relationships
      ii. Example: Figure 11-6: Once solubility is reached, no more can dissolve
4. Dissolving process is a reversible reaction
   a. ***Unsaturated*** – less than standard amount of solute – no undissolved solute remains
   b. ***Saturated*** – standard amount of solute – some solute remains undissolved
      i. Example: Figure 11-8: High humidity = saturated air = slow evaporation = "sticky" feeling when it's hot
   c. ***Supersaturated*** – contains more than standard amount of solute
      i. Example: rain, fog, dew, frost, other precipitation – comes out of solution because there is too much!
5. Some substances do not dissolve
   a. Some metals (copper, zinc, iron) and salts are insoluble in water

Name ________________________ Date ____________ Class ____________

CHAPTER 12 REVIEW

Solutions

SECTION 1

SHORT ANSWER Answer the following questions in the space provided.

1. Match the type of mixture on the left to its representative particle diameter on the right.

_____ solutions (a) larger than 1000 nm

_____ suspensions (b) 1 nm to 1000 nm

_____ colloids (c) smaller than 1 nm

2. Identify the solvent in each of the following examples:

________________ **a.** tincture of iodine (iodine dissolved in ethyl alcohol)

________________ **b.** sea water

________________ **c.** water-absorbing super gels

3. A certain mixture has the following properties:

- No solid settles out during a 48-hour period.
- The path of a flashlight beam is easily seen through the mixture.
- It appears to be homogeneous under a hand lens but not under a microscope.

Is the mixture a suspension, colloid, or true solution? Explain your answer.

____

____

____

____

4. Define each of the following terms:

a. alloy

____

____

b. electrolyte

____

____

c. aerosol

____

____

d. aqueous solution

____

____

5. For each of the following types of solutions, give an example other than those listed in **Table 1** on page 402 of the text:

a. a gas in a liquid

____

____

b. a liquid in a liquid

____

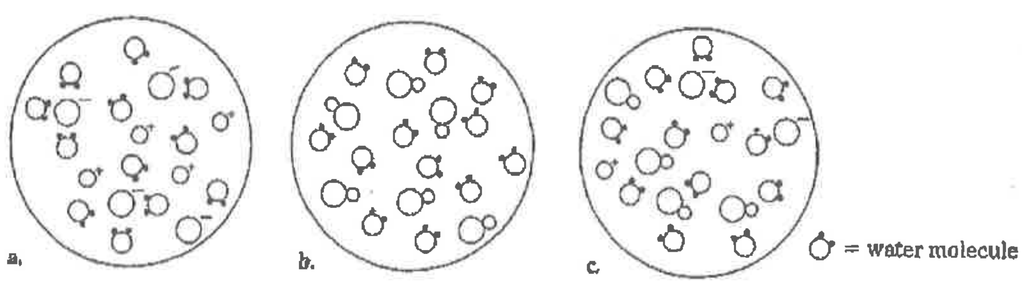
____

c. a solid in a liquid

____

____

6. Using the following models of solutions shown at the particle level, indicate which will conduct electricity. Give a reason for each model.



a. ____

b. ____

c. ____

Name ______________________ Date ____________ Class ____________

CHAPTER 12 REVIEW

Solutions

SECTION 2

SHORT ANSWER **Answer the following questions in the space provided.**

1. The following are statements about the dissolving process. Explain each one at the molecular level.

a. Increasing the pressure of a solute gas above a liquid solution increases the solubility of the gas in the liquid.

____

____

____

____

b. Increasing the temperature of water speeds up the rate at which many solids dissolve in this solvent.

____

____

____

____

____

____

c. Increasing the surface area of a solid solute speeds up the rate at which it dissolves in a liquid solvent.

____

____

____

____

2. The solubility of $KClO_3$ at 25°C is 10. g of solute per 100. g of H_2O.

a. If 15 g of $KClO_3$ are stirred into 100 g of water at 25°C, how much of the $KClO_3$ will dissolve? Is the solution saturated, unsaturated, or supersaturated?

____

____

b. If 15 g of $KClO_3$ are stirred into 200 g of water at 25°C, how much of the $KClO_3$ will dissolve? Is the solution saturated, unsaturated, or supersaturated?

____

____

PROBLEMS **Write the answer on the line to the left. Show all your work in the space provided.**

3. Use the data in **Table 4** on page 410 of the text to answer the following questions:

____________ **a.** How many grams of LiCl are needed to make a saturated solution with 300. g of water at 20°C?

____________ **b.** What is the minimum amount of water needed to dissolve 51 g of $NaNO_3$ at 40°C?

____________ **c.** Which solute forms a saturated solution when 36 g of it are dissolved in 25 g of water at 20°C?

4. KOH is an ionic solid readily soluble in water.

____________ **a.** What is its enthalpy of solution in kJ/g? Refer to the data in **Table 5** on page 416 of the text.

b. Will the temperature of the system increase or decrease as the dissolution of KOH proceeds? Why?

____

____

____

Name ____________________ Date ____________ Class ____________

CHAPTER 12 REVIEW

Solutions

SECTION 3

SHORT ANSWER Answer the following questions in the space provided.

1. Describe the errors made by the following students in making molar solutions.

a. James needs a 0.600 M solution of KCl. He measures out 0.600 g of KCl and adds 1 L of water to the solid.

____

____

____

____

b. Mary needs a 0.02 M solution of $NaNO_3$. She calculates that she needs 2.00 g of $NaNO_3$ for 0.02 mol. She puts this solid into a 1.00 L volumetric flask and fills the flask to the 1.00 L mark.

____

____

____

____

PROBLEMS Write the answer on the line to the left. Show all of your work in the space provided.

2. ____________ What is the molarity of a solution made by dissolving 2.0 mol of solute in 6.0 L of solvent?

3. ____________ CH_3OH is soluble in water. What is the molality of a solution made by dissolving 8.0 g of CH_3OH in 250. g of water?

4. Marble chips effervesce when treated with hydrochloric acid. This reaction is represented by the following equation:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

To produce a reaction, 25.0 mL of 4.0 M HCl is added to excess $CaCO_3$.

______________ **a.** How many moles of HCl are consumed in this reaction?

______________ **b.** How many liters of CO_2 are produced at STP?

______________ **c.** How many grams of $CaCO_3$ are consumed?

5. Tincture of iodine is $I_2(s)$ dissolved in ethanol, C_2H_5OH. A 1% solution of tincture of iodine is 10.0 g of solute for 1000. g of solution.

______________ **a.** How many grams of solvent are present in 1000. g of this solution?

______________ **b.** How many moles of solute are in 10.0 g of I_2?

______________ **c.** What is the molality of this 1% solution?

d. To determine a solution's molarity, the density of that solution can be used. Explain how you would use the density of the tincture of iodine solution to calculate its molarity.

____

____

____

____

Name ______________________ Date ______________ Class ______________

CHAPTER 12 REVIEW

Solutions

MIXED REVIEW

SHORT ANSWER Answer the following questions in the space provided.

1. Solid $CaCl_2$ does not conduct electricity. Explain why it is considered to be an electrolyte.

____

____

____

2. Explain the following statements at the molecular level:

a. Generally, a polar liquid and a nonpolar liquid are immiscible.

____

____

____

b. Carbonated soft drinks taste flat when they warm up.

____

____

____

____

3. An unknown compound is observed to mix with toluene, $C_6H_5CH_3$, but not with water.

a. Is the unknown compound ionic, polar covalent, or nonpolar covalent? Explain your answer.

____

____

b. Suppose the unknown compound is also a liquid. Will it be able to dissolve table salt? Explain why or why not.

____

____

____

PROBLEMS **Write the answer on the line to the left. Show all your work in the space provided.**

4. Consider 500. mL of a 0.30 M $CuSO_4$ solution.

____________ **a.** How many moles of solute are present in this solution?

____________ **b.** How many grams of solute were used to prepare this solution?

5. a. If a solution is electrically neutral, can all of its ions have the same type of charge? Explain your answer.

____

____

____________ **b.** The concentration of the OH^- ions in pure water is known to be 1.0×10^{-7} M. How many OH^- ions are present in each milliliter of pure water?

6. 90. g of $CaBr_2$ are dissolved in 900. g of water.

____________ **a.** What volume does the 900. g of water occupy if its density is 1.00 g/mL?

____________ **b.** What is the molality of this solution?

Name ______________________________ Date________________

Class Notes: Chapter 13 Ions in Aqueous Solutions and Colligative Properties

Dissociation

The separation of ions when an ionic compound dissolves is dissociation.

$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$ 2 ions formed

$CaCl_2(s) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$ 3 ions formed

Write the equation for the dissolution of 1 mole of Ammonium chloride.

$NH_4Cl(s) \rightarrow NH_4^+(aq) + Cl^-(aq)$

One mole = two moles of ions

Write the equation for the dissolution of 2 moles of
Sodium sulfide. (Na_2S)

The Solubility Rules

1. All common salts of the Group 1A elements and ammonium are soluble.
2. All common acetates and nitrates are soluble.
3. All binary compounds of Group VIIA elements (other than F) with metals are soluble except those of silver, mercury (I), and lead.
4. All sulfates are soluble except those of barium, strontium, lead, calcium, silver, and mercury(I).
5. Except for those in Rule 1, carbonates, hydroxides, oxides, and phosphates are insoluble.

Write a summary, a complete, and a net ionic equation for the reaction between the aqueous solutions, $Cd(NO_3)$ and $(NH_4)_2S$ Note: A precipitate does form in this reaction.(Refer to text: page 439)

Will a precipitate form if solutions of potassium nitrate and magnesium sulfate are combined? If so, write the net ionic equation.

Will a precipitate form if solutions of barium chloride and sodium sulfate are combined? If so, write the net ionic equation and identify the spectator ions.

Copper(II)chloride and lead(II)nitrate react in aqueous solutions by double displacement. Write the summary equation, the complete ionic equation, and the net ionic equation. Circle the spectators in the complete ionic equation.

Ionization is a process of forming ions from molecular substances.

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$ This is the hydronium ion.
When lots of ions are formed, the solution is a strong electrolytic solution. This means it conducts electricity very well.

Weak electrolytic solutions are a combination of molecules and ions.

$HC_2H_3O_2 + H_2O \rightleftarrows H_3O^+ + C_2H_3O_2^-$

The double arrow means the reaction is reversible. At equilibrium, there are more molecules than ions present. This solution will conduct electricity poorly.

Colligative Properties
Definition:
The properties of a solution that are dependent only on the number of solute particles in solution (not the identity of the solute).
- Vapor pressure lowering
- Boiling point elevation
- Freezing point depression

•Osmotic pressure

A **nonvolatile** solute does not have the tendency to become a gas at room conditions.

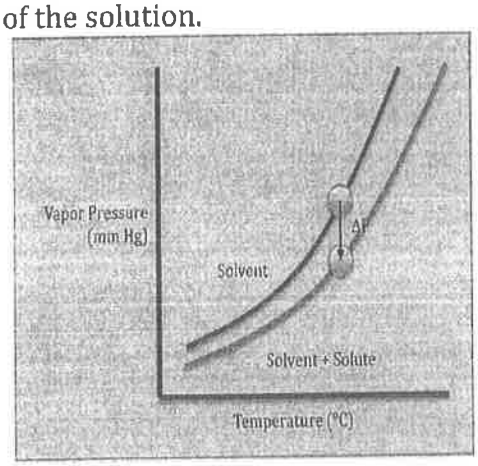
A **volatile** solute DOES have the tendency to become a gas at room conditions.

According to <u>**Raoult's**</u> <u>**Law**</u>, the vapor pressure of a solution will be lower when a solute is present. The more particles present, the lower the vapor pressure.

When the solute-solvent attractions are stronger than the original solvent-solvent attractions, the solution vapor pressure will be lower. This is typically the case when the solute is nonvolatile.

Definition:
When the vapor pressure of a solvent containing a solute is compared to the vapor pressure of the pure solvent, the solution's vapor pressure is found to be lower.
$P_{solv}^{\circ} - P_{solv} = \Delta P$

¨In the diagram notice that the vapor pressure lowers when a solvent is mixed with a nonvolatile solute.
¨The P_{solv}° increases as the temperature increases; similarly, the P_{solv} increases as the temperature increases.
¨This means that the lowering of the vapor pressure leads to a higher boiling point of the solution.



Definition:
The boiling point of a solution is greater than the boiling point of the pure solvent because the solution (which has a *lower vapor pressure) will need to be heated to a higher temperature in order for the vapor pressure to become equal to the external pressure (i.e., the boiling point).*

Boiling occurs when the vapor pressure above the solution = the atmospheric pressure.

The boiling point elevation, ΔT_b, is the difference between the higher boiling point of the solution and the boiling point of the pure solvent.

$\Delta T_b = T_b - T_b^\circ$ i

If the solute is molecular, then (i) = 1. Molecular substances do not break apart into more particles in solution.

If the solute is ionic, then (i) depends on the formula. Example: NaCl creates 2 particles in solution. Na^+ and Cl^-.

When a nonelectrolyte solute dissolves in a solvent, it dissolves without separating into ions.

Like methanol dissolving in water...

$CH_3OH(l) \rightarrow CH_3OH(aq)$

When an electrolyte solute dissolves in a solvent, it does separate into ions.

Like sodium sulfate dissolving in water...

$Na_2SO_4(s) \rightarrow 2\ Na^+(aq) + 1\ SO_4^{2-}(aq)$

Since colligative properties depend on the number of solute particles, not the identity of the solute, an electrolyte influences those properties more than a nonelectrolyte.

A 1.0 M sodium sulfate solution creates how many particles?

$Na_2SO_4(s)\quad 2\ Na^+(aq) + 1\ SO_4^{2-}(aq)$

A strong electrolyte, such as Na_2SO_4, will dissolve 100 %, creating one particle for each dissociated ion.

These particles are called the van't Hoff factor, *i*.

Not all ionic substances influence the boiling/freezing point.

A weak electrolyte, such as Ag_3PO_4, will NOT dissolve 100 %. This means that fewer particles will be able to influence the colligative properties.

How many particles would the following solutes provide?

$CaCl_2$

$Mg(NO_3)_2$

$C_6H_{12}O_6$

$Al_2(SO_4)_3$

The boiling point elevation, ΔT_b, is directly proportional to the molality of the solute.

$\Delta T_b = k_b \bullet m_{solute}\, i$

"Each solvent has a different boiling point constant.

Solvent	T_b° of pure solvent (°C)	k_b (°C/m)
Water	100.00	+0.5121
Benzene	80.10	+2.53
Camphor	207.4	+5.611
Chloroform	61.70	+3.63

$\Delta T_b = k_b \bullet m_{solute}\, i$

A solution was made up of eugenol, $C_{10}H_{12}O_2$, in diethyl ether. If the solution was 0.575 m eugenol in ether, what was the boiling point of the solution? The boiling point of pure diethylether is 34.6°C and the boiling-point-elevation constant is 2.02 °C/m.

Addition of a nonvolatile solute to the volatile solvent increases the attractions (and lowers the vapor pressure). This means that the particles are closer together so a lower temperature allows them to freeze.

The freezing point depression, ΔT_f, is the difference between the higher freezing point of the pure solvent and the freezing point of the solution.

$\Delta T_f = T_f^\circ - T_f$

The freezing point depression, ΔT_f, is directly proportional to the molality of the solute.

$\Delta T_f = k_f \bullet m_{solute}\, i$

"Each solvent has a different freezing point constant.

Solvent	T_f° of pure solvent (°C)	k_b (°C/m)
Water	0.0	+1.86
Benzene	5.50	+5.12
Camphor	179.95	+39.7

$\Delta T_f = k_f \bullet m_{solute}\, i$

A solution was made up of 0.575 m eugenol, $C_{10}H_{12}O_2$, in diethyl ether. What was the freezing point of the solution? The freezing point of diethyl ether is -116.3°C and the freezing-point-depression constant is 1.79 °C/m.

Osmosis is the diffusion of small molecules through a semi-permeable membrane. Usually, osmosis is seen in the net movement of the solvent from the pure solvent (low solute concentration) to solution (high solute concentration).

The membrane is termed "semi-permeable" because small molecules such as water or small ions (Na^+ or K^+) may pass in either direction through the membrane. "Osmotic pressure is the pressure necessary to just stop osmosis. This is done by pressing on the solution side to increase the movement of solvent particles from the solution back into the pure solvent.

Problems:

1.What is the boiling point elevation when 11.4 g of ammonia is dissolved in 200.g of water?

2. How many grams of benzoic acid ($C_7H_6O_2$) must be dissolved in 78.1g of ethanol to raise the boiling point by 4.00 Celsius degrees. Kb for ethanol is 1.20 C°/m

3. What is the boiling point elevation when 10.0 g of magnesium nitrate is added to 500. grams of water?

4. How many grams of sodium chloride must be dissolved in 250.g of water to raise the boiling point by 2.50 Celsius degrees?

5. How many grams of silver would have to be dissolved in 1120 grams of ethanol to lower the freezing point by 0.25 Celsius degrees? K_f for ethanol is 1.99C°/m

6. What is the freezing point depression when 85.3 grams of oxygen is dissolved in 1500 grams of water?

7. How many grams of calcium chloride are needed to lower the freezing point of 2100.g of water by 5.0 Celsius degrees?

8. If you lower the freezing point of 16.8 grams of chloroform by 2.50 Celsius degrees by using chlorine gas, how many grams of chlorine gas must be dissolved in the chloroform? K_f for chloroform is 4.68 C°/m

Name ______________ Date ______________

Chap 13 Day 2

Colligative Properties

Definition:

The properties of a solution that are dependent only on the number of solute particles in solution (not the identity of the solute).

- Vapor pressure lowering
- Boiling point elevation
- Freezing point depression
- Osmotic pressure

A **nonvolatile** solute does not have the tendency to become a gas at room conditions.

A **volatile** solute DOES have the tendency to become a gas at room conditions.

According to **Raoult's Law**, the vapor pressure of a solution will be lower when a solute is present. The more particles present, the lower the vapor pressure.

When the solute-solvent attractions are stronger than the original solvent-solvent attractions, the solution vapor pressure will be lower. This is typically the case when the solute is nonvolatile.

Definition:

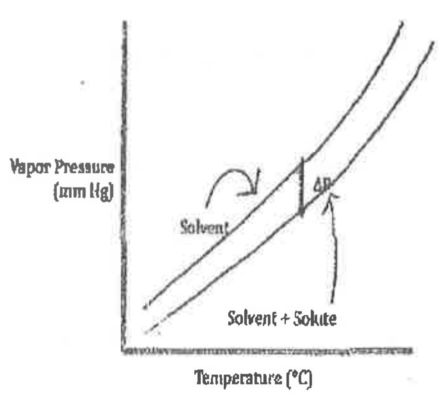
When the vapor pressure of a solvent containing a solute is compared to the vapor pressure of the pure solvent, the solution's vapor pressure is found to be lower.

$P_{solv}° - P_{solv} = \Delta P$

"In the diagram notice that the vapor pressure lowers when a solvent is mixed with a nonvolatile solute.

"The $P_{solv}°$ increases as the temperature increases; similarly, the P_{solv} increases as the temperature increases.

"This means that the lowering of the vapor pressure leads to a higher boiling point of the solution.



Solvent	Formula	Freezing Point (°C)	K_f (°C/molal)	Boiling Point (°C)	K_b (°C/molal)
Water	H_2O	0.0	-1.86	100.0	0.51
Acetic acid	CH_3COOH	17.0	-3.90	118.1	3.07
Benzene	C_6H_6	5.5	-4.90	80.2	2.53
Chloroform	$CHCl_3$	-63.5	-4.68	61.2	3.63
Ethanol	C_2H_5OH	-114.7	-1.99	78.4	1.22
Phenol	C_6H_5OH	43.0	-7.40	181.0	3.56

Table 1. Molal Freezing Point and Boiling Point Constants

Definition:

The boiling point of a solution is greater than the boiling point of the pure solvent because the solution (which has a *lower vapor pressure) will need to be heated to a higher temperature in order for the vapor pressure to become equal to the external pressure (i.e., the boiling point).*

Boiling occurs when the vapor pressure above the solution = the atmospheric pressure.

The boiling point elevation, ΔT_b, is the difference between the higher boiling point of the solution and the boiling point of the pure solvent.

$\Delta T_b = T_b - T_b^\circ\ i$

If the solute is molecular, then (i) = 1. Molecular substances do not break apart into more particles in solution.

If the solute is ionic, then (i) depends on the formula. Example: NaCl creates 2 particles in solution. Na^+ and Cl^-.

When a nonelectrolyte solute dissolves in a solvent, it dissolves without separating into ions.

Like methanol dissolving in water...

$CH_3OH(l) \rightarrow CH_3OH(aq)$

When an electrolyte solute dissolves in a solvent, it does separate into ions.

Like sodium sulfate dissolving in water...
$Na_2SO_4(s) \rightarrow$ 2 $Na^+(aq)$ + 1 $SO_4^{2-}(aq)$

Since colligative properties depend on the number of solute particles, not the identity of the solute, an electrolyte influences those properties more than a nonelectrolyte.

A 1.0 M sodium sulfate solution creates how many particles?
$Na_2SO_4(s)$ 2 $Na^+(aq)$ + 1 $SO_4^{2-}(aq)$

A strong electrolyte, such as Na_2SO_4, will dissolve 100 %, creating one particle for each dissociated ion.

These particles are called the van't Hoff factor, *i*.

Not all ionic substances influence the boiling/freezing point.

A weak electrolyte, such as Ag_3PO_4, will NOT dissolve 100 %. This means that fewer particles will be able to influence the colligative properties.

How many particles would the following solutes provide?
$CaCl_2$
$Mg(NO_3)_2$
$C_6H_{12}O_6$
$Al_2(SO_4)_3$
The boiling point elevation, ΔT_b, is directly proportional to the molality of the solute.

$\Delta T_b = k_b \bullet m_{solute}\, i$
Each solvent has a different boiling point constant.

$\Delta T_b = k_b \bullet m_{solute}\, i$
A solution was made up of eugenol, $C_{10}H_{12}O_2$, in diethyl ether. If the solution was 0.575 m eugenol in ether, what was the boiling point of the solution? The boiling point of pure diethylether is 34.6°C and the boiling-point-elevation constant is 2.02 °C/m.

Addition of a nonvolatile solute to the volatile solvent increases the attractions (and lowers the vapor pressure). This means that the particles are closer together so a lower temperature allows them to freeze.

The freezing point depression, ΔT_f, is the difference between the higher freezing point of the *pure solvent* and the freezing point of the solution.

$\Delta T_f = T_f^o - T_f$

The freezing point depression, ΔT_f, is directly proportional to the molality of the solute.

$\Delta T_f = k_f \bullet m_{solute}\, i$

Each solvent has a different freezing point constant.

$$\Delta T_f = k_f \bullet m_{solute}\, i$$

A solution was made up of 0.575 m eugenol, $C_{10}H_{12}O_2$, in diethyl ether. What was the freezing point of the solution? The freezing point of diethyl ether is -116.3°C and the freezing-point-depression constant is 1.79 °C/m.

Osmosis is the diffusion of small molecules through a semi-permeable membrane. Usually, osmosis is seen in the net movement of the solvent from the pure solvent (low solute concentration) to solution (high solute concentration).

The membrane is termed "semi-permeable" because small molecules such as water or small ions (Na^+ or K^+) may pass in either direction through the membrane.

Osmotic pressure is the pressure necessary to just stop osmosis. This is done by pressing on the solution side to increase the movement of solvent particles from the solution back into the pure solvent

Problems:

1.What is the boiling point elevation when 11.4 g of ammonia is dissolved in 200.g of water?

2. How many grams of benzoic acid ($C_7H_6O_2$) must be dissolved in 78.1g of ethanol to raise the boiling point by 4.00 Celsius degrees. Kb for ethanol is 1.20 C°/*m*

3. What is the boiling point elevation when 10.0 g of magnesium nitrate is added to 500. grams of water?

4. How many grams of sodium chloride must be dissolved in 250.g of water to raise the boiling point by 2.50 Celsius degrees?

5. How many grams of silver would have to be dissolved in 1120 grams of ethanol to lower the freezing point by 0.25 Celsius degrees? K_f for ethanol is 1.99C°/*m*

6. What is the freezing point depression when 85.3 grams of oxygen is dissolved in 1500 grams of water?

7. How many grams of calcium chloride are needed to lower the freezing point of 2100.g of water by 5.0 Celsius degrees?

8. If you lower the freezing point of 16.8 grams of chloroform by 2.50 Celsius degrees by using chlorine gas, how many grams of chlorine gas must be dissolved in the chloroform? K_f for chloroform is 4.68 C°/*m*

Name ______________________ Date ____________ Class ____________

CHAPTER 13 REVIEW

Ions in Aqueous Solutions and Colligative Properties

SECTION 1

SHORT ANSWER Answer the following questions in the space provided.

1. Use the guidelines in **Table 1** on page 437 of the text to predict the solubility of the following compounds in water:

______________ **a.** magnesium nitrate

______________ **b.** barium sulfate

______________ **c.** calcium carbonate

______________ **d.** ammonium phosphate

2. 1.0 mol of magnesium acetate is dissolved in water.

______________ **a.** Write the formula for magnesium acetate.

______________ **b.** How many moles of ions are released into solution?

______________ **c.** How many moles of ions are released into a solution made from 0.20 mol magnesium acetate dissolved in water?

3. Write the formula for the precipitate formed

______________ **a.** when solutions of magnesium chloride and potassium phosphate are combined.

______________ **b.** when solutions of sodium sulfide and silver nitrate are combined.

4. Write ionic equations for the dissolution of the following compounds:

a. $Na_3PO_4(s)$

____

b. iron(III) sulfate(*s*)

____

5. **a.** Write the net ionic equation for the reaction that occurs when solutions of lead(II) nitrate and ammonium sulfate are combined.

____

b. What are the spectator ions in this system?

____

SECTION 1 continued

6. The following solutions are combined in a beaker: NaCl, Na_3PO_4, and $Ba(NO_3)_2$.

a. Will a precipitate form when the above solutions are combined? If so, write the name and formula of the precipitate.

____

____

b. List all spectator ions present in this system.

____

____

7. It is possible to have spectator ions present in many chemical systems, not just in precipitation reactions. Consider this example:

$$Al(s) + HCl(aq) \rightarrow AlCl_3(aq) + H_2(g) \text{ (unbalanced)}$$

________________ **a.** In an aqueous solution of HCl, virtually every HCl molecule is ionized. True or False?

________________ **b.** There is only one spectator ion in this system. Is it $Al^{3+}(aq)$, $H^+(aq)$, or $Cl^-(aq)$?

c. Balance the above equation.

____

________________ **d.** If 9.0 g of Al metal react with excess HCl according to the balanced equation in part **c**, what volume of hydrogen gas at STP will be produced? Show all your work.

8. Acetic acid, CH_3CO_2H, is a weak electrolyte. Write an equation to represent its ionization in water. Include the hydronium ion, H_3O^+.

____

Name ____________________ Date __________ Class __________

CHAPTER 13 REVIEW

Ions in Aqueous Solutions and Colligative Properties

SECTION 2

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

1. ____________ **a.** Predict the boiling point of a 0.200 *m* solution of glucose in water.

____________ **b.** Predict the boiling point of a 0.200 *m* solution of potassium iodide in water.

2. A chief ingredient of antifreeze is liquid ethylene glycol, $C_2H_4(OH)_2$. Assume $C_2H_4(OH)_2$ is added to a car radiator that is holding 5.0 kg of water.

____________ **a.** How many moles of ethylene glycol should be added to the radiator to lower the freezing point of the water from 0°C to −18°C?

____________ **b.** How many grams of ethylene glycol does the quantity in part a represent?

____________ **c.** Ethylene glycol has a density of 1.1 kg/L. How many liters of $C_2H_4(OH)_2$ should be added to the water in the radiator to prevent freezing down to −18°C?

d. In World War II, soldiers in the Sahara Desert needed a supply of antifreeze to protect the radiators of their vehicles. The temperature in the Sahara almost never drops to 0°C, so why was the antifreeze necessary?

____

____

____

____

3. An important use of colligative properties is to determine the molar mass of unknown substances. The following situation is an example: 12.0 g of unknown compound X, a nonpolar nonelectrolyte, is dissolved in 100.0 g of melted camphor. The resulting solution freezes at 99.4°C. Consult **Table 2** on page 448 of the text for any other data needed to answer the following questions:

______________ **a.** By how many °C did the freezing point of camphor change from its normal freezing point?

______________ **b.** What is the molality of the solution of camphor and compound X, based on freezing-point data?

______________ **c.** If there are 12.0 g of compound X per 100.0 g of camphor, how many grams of compound X are there per kilogram of camphor?

______________ **d.** What is the molar mass of compound X?

4. Explain why the ability of a solution to conduct an electric current is not a colligative property.

____

____

____

Name ____________________ Date ____________ Class ____________

CHAPTER 13 REVIEW

Ions in Aqueous Solutions and Colligative Properties

MIXED REVIEW

SHORT ANSWER Answer the following questions in the space provided.

1. Match the four compounds on the right to their descriptions on the left.

______	an ionic compound that is quite soluble in water	(a) HCl
______	an ionic compound that is not very soluble in water	(b) $NaNO_3$
______	a molecular compound that ionizes in water	(c) AgCl
______	a molecular compound that does not ionize in water	(d) C_2H_5OH

2. Consider nonvolatile nonelectrolytes dissolved in various liquid solvents to complete the following statements:

____________________ **a.** The change in the boiling point does *not* vary with the identity of the ____ (solute, solvent), assuming all other factors remain constant.

____________________ **b.** The change in the boiling point varies with the identity of the ____ (solute, solvent), assuming all other factors remain constant.

____________________ **c.** The change in the boiling point becomes greater as the concentration of the solute in solution ____ (increases, decreases).

3. **a.** Name two compounds in solution that could be combined to cause the formation of a calcium carbonate precipitate.

____

____

b. Identify any spectator ions in the system you described in part **a.**

____

c. Write the net ionic equation for the formation of calcium carbonate.

____

4. Explain why applying rock salt (impure NaCl) to an icy sidewalk hastens the melting process.

____

____

____

MIXED REVIEW continued

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

5. ______________ Some insects survive cold winters by generating an antifreeze inside their cells. The antifreeze produced is glycerol, $C_3H_5(OH)_3$, a nonvolatile nonelectrolyte that is quite soluble in water. What must the molality of a glycerol solution be to lower the freezing point of water to −25.0°C?

6. ______________ How many grams of methanol, CH_3OH, should be added to 200. g of acetic acid to lower its freezing point by 1.30°C? Refer to **Table 2** on page 448 of the text for any necessary data.

7. ______________ The boiling point of a solution of glucose, $C_6H_{12}O_6$, and water was recorded to be 100.34°C. Calculate the molality of this solution.

8. HF(*aq*) is a weak acid. A 0.05 mol sample of HF is added to 1.0 kg of water.

a. Write the equation for the ionization of HF to form hydronium ions.

____

______________________ **b.** If HF became 100% ionized, how many moles of its ions would be released?

9. ______ Which solution has the highest osmotic pressure?

a. 0.1 *m* glucose
b. 0.1 *m* sucrose
c. 0.5 *m* glucose
d. 0.2 *m* sucrose

Name____ Period__________

Test Review

1. Predict the products of the following reaction and note the solid. (Use the Solubility Table) Balance the equation.

   $CsOH + AlCl_3 \rightarrow$

   Ionic Equation. ____

   Net Ionic Equation ____

   List the spectator ions: ____

2. Predict the products of the following reaction and note the solid. (Use the Solubility Table) Balance the equation.

   $HC_2H_3O_2 + AgNO_3 \rightarrow$

   Ionic Equation. ____

   Net Ionic Equation ____

   List the spectator ions: ____

3. Briefly describe how a nonvolatile solute affects each of the following.
   a. Vapor pressure
   b. Freezing point
   c. Boiling point
   d. Osmotic pressure
4. What is the new boiling point of a 2.5 m solution of $C_6H_{12}O_6$? ($K_b = 0.512\ ^\circ C/m$)

5. What is the new freezing point of a solution MgI_2, when 556.6 g is dissolved in 2.00 kg of water? ($K_f = 1.86\ ^\circ C/m$)

6. A molecular compound having a mass of 92.0 g was dissolved in 1000.0 g of water. The freezing point of the solution was lowered to -3.72°C. Determine the molecular mass of this compound. ($K_f = 1.86\ ^\circ C/m$)

7. Write the ionization equation for HNO_3 and water.

8. Write the ionization equation for $HC_2H_3O_2$ and water

9. The boiling point of naphthalene is 217.2°C. If 120.0 g of a molecular solute is added to 2.0 kg of naphthalene and the new boiling point is 223.50°C, what is the molecular mass of the solute? (K_b = 5.80 °C/m)

10. Vocabulary to know: dissociation, net ionic equation, spectator ions, ionization, hydronium ion, strong electrolyte, weak electrolyte, colligative properties, nonvolatile substance, molal freezing-point constant, freezing-point depression, molal boiling-point constant, boiling-point elevation, semipermeable membrane, osmosis, osmotic pressure.

Name ____ 2-24-14

Practice Test for Chapters 12 and 13

1. Draw atomic level pictures of a saturated solution, a colloid, a unsaturated solution, a suspension.

2. Define the following terms:

   Colloid

   Solution

   Alloy

   Tyndall effect

   Electrolyte

   Nonelectrolyte

3. Give an example of an electrolyte and a nonelectrolyte.

4. How can you tell if something is a homogeneous solution?

5. Write the formula for the hydronium ion. Why does a hydronium ion form?

6. What can increase the dissolving rate of a solid solute?

7. What can make a gaseous solute come out of solution?

8. How is the solubility of a substance determined?

9. Which type of solution forms crystals if disturbed?

10. Explain the enthalpy of an endothermic dissolution. (What might you feel and how is this expressed numerically?)

11. Explain the enthalpy of an exothermic dissolution. (What might you feel and how is this expressed numerically?)

12. Give examples of "like dissolving like" .

13. Name a common polar solvent.

14. Name a common nonpolar solvent.

15. What happens during the solvation process?

16. State Henry's Law.

17. What volume of 6.0M sulfuric acid do you need to create 100.0mL of a 2.0M solution?

18. What is the molarity of a solution that contains 4.0g of NaCl in 400mL of solution?

19. How many grams of NaCl would you need to make a 3.00m solution with 500.g of water?

20. How many moles of ions are produced when one mole of $Mg(NO_3)_2$ dissociates.

21. Write the dissociation equation for magnesium nitrate.

22. Write the general, complete and net ionic equation for the reaction between two solutions of lead (II) chloride and aluminum sulfate.

23. What is the difference between ionization and dissociation?

24. What is the definition of nonvolatile? How does a nonvolatile solute affect the F.P. and B.P. of a pure solvent?

25. What is the change in the boiling point of a 1.25m solution of $PbCl_2$ as compared to the pure solvent (water.)

26. 300.g of a nonelectrolyte is dissolved in 250.0mL of water. The boiling point of the solution is measured to be 102.7oC. What is the molar mass of this compound? Kb is 0.51°C/m

Study guide for Chapters 12 and 13

Know the definitions of the following terms:

Dilute	Concentrated	Solute	Solvent
Molarity	Molality	Colligative	Vapor pressure
Alloy	Solution	Emulsion	Suspension
Electrolyte	nonelectrolyte	surface area	Tyndall effect
Colloid	solvation	solubility	saturated
Supersaturated	unsaturated	polar	nonpolar
Henry's Law	precipitation	homogeneous	heterogeneous

1. Be able to identify homogeneous and heterogeneous mixtures.
2. Know that solutions have particles that are dissolved in them and that dissolved particles are extremely tiny (atomic level tiny).
3. Know how to determine if a mixture is a true solution or a colloid.
4. Know the formula for a hydronium ion.
5. Be ready to solve for boiling point elevation and freezing point depression.
6. Know what a nonvolatile solute does in regards to freezing and boiling points of pure solvents.
7. Be able to identify electrolytes from nonelectrolytes.
8. Know factors that increase the rate of dissolving of a solid.
9. Know factors that increase the solubility of a gas.
10. Know what happens during the process of solvation.
11. Know how to read a solubility graph.
12. Know what "Like dissolves like " means.
13. Know the sign of the enthalpy value for an exothermic reaction and an endothermic reaction.
14. Be able to identify polar and nonpolar solvents.
15. Be ready to solve Molarity, molality and dilution problems.
16. Know how to figure out the number of ions if given a formula.
17. Know how to write complete ionic and net ionic equations.
18. Know how to read solubility rules.
19. Know how to predict a precipitate using solubility rules.
20. Be able to do a stoichiometry problem and solve for mL of solution needed. (Solve for moles and then use the Molarity formula)
21. Be able to write dissociation equations for a salt when given the formula.
22. Be able to use the elevated boiling point, the Kb, the grams and the amount of kg of solvent to calculate the molar mass of the compound. (Solve for moles and then divide grams by moles)