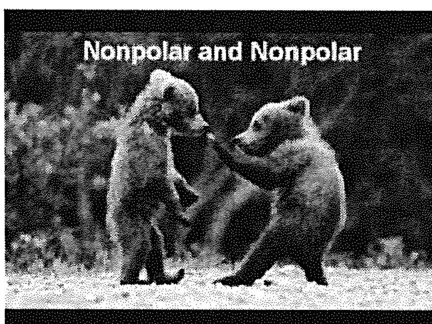
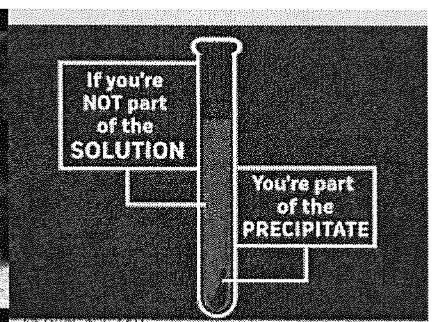
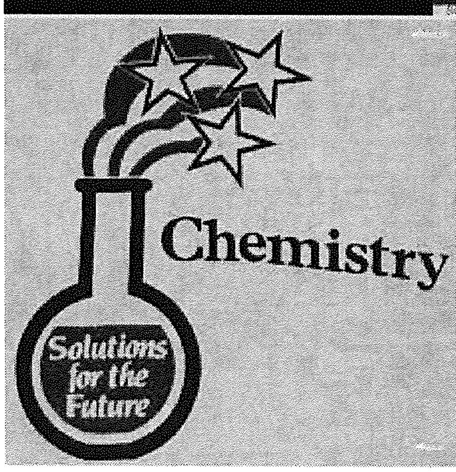
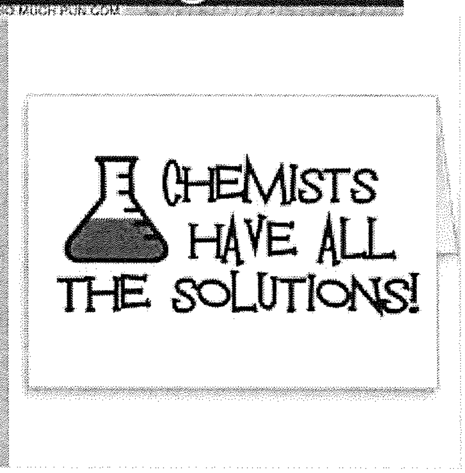
Chapter 13 – Solution Chemistry

What are solutions?
Types of solutions and how they are made.
Solubility Curves – Table G
Molarity
Molality
Electrolytes
Molecular, ionic, and net ionic equations
Colligative properties
Calculating the effect of a solute on freezing and boiling points









Name ____________________

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Solutions: An Introduction

Aim

- To explain why substances dissolve

Notes

Definition: Solution = homogeneous mixture

- ★ Nature of mixtures
  - ☆ consists of two or more kinds of matter
  - ☆ each substance in a mixture retains its own properties
    - ✯ sugar and water - sweet and wet
    - ✯ brine (salt water) - salty liquid
  - ☆ the composition is variable (not constant)
  - ☆ can be separated by physical means
- ★ Distinguishing solutions from mechanical mixtures
  - ☆ properties of solutions
    - ✯ homogeneous mixtures – composed of two or more substances and have variable composition *BUT* the particles are distributed evenly throughout each other *SO* the composition is uniform
      - ✯ the solution appears to be one substance
    - ✯ consist of a ***solute*** dissolved in a ***solvent***
  - ☆ solute - substance that *IS* dissolved by another
  - ☆ solvent
    - ✯ substance that dissolves another
    - ✯ continuous phase - salt dissolved in water appears to be a liquid

Solubility - ability to dissolve in water

- ★ Factors that affect solubility
  - ☆ Degree of solubility (how much dissolves)
    - ✯ nature of solute and solvent
      - ✯ in order for a solvent to dissolve a solute, it must exert forces of attraction on the solute
        - ✯ polar solvents such as water dissolve polar and ionic solutes because they exert mutual attractions that cause their particles to intermingle
        - ✯ nonpolar solvents such as benzene do NOT dissolve polar and ionic substances because they exert no forces of attraction that would cause the particles to separate so they can intermingle
          - ★ oil and water do NOT mix
        - ✯ nonpolar substances such as fat dissolve in nonpolar solvents such as benzene because the forces of attraction are too weak to prevent the particles from freely intermingling
      - ✯ like dissolves like (See Table F - Table of Solubilities in Water)
    - ✯ Temperature (See Table G - Solubility Curves)
      - ✯ solubility of solid solutes generally increases as temperature increases
      - ✯ solubility of gaseous solutes generally decreases as temperature increases
    - ✯ Pressure
      - ✯ solids and liquids - no effect
      - ✯ gases: Henry's Law - mass of a dissolved gas in a liquid is directly proportional to the pressure of the gas
  - ☆ Rate of solution

Factor	Affect on Solid Solute	Affect on Gaseous Solute
Particle Size	reducing particle size by crushing increases the rate by increasing surface area	not applicable
Stirring	increases the rate by exposing fresh solvent to solute and increasing kinetic energy	decreases the rate by increasing kinetic energy, thereby reducing solubility
Amount of dissolved solute	as the amount of dissolved solute increases, the rate decreases	as the amount of dissolved solute increases, the rate decreases
Temperature	as the temperature increases, the rate increases	as the temperature increases, the rate decreases

- ★ Saturation (see Table G)
  - ☆ Saturated solution - solution that cannot dissolve any more solute at a given temperature
    - ✯ added solute will NOT dissolve
  - ☆ Unsaturated solution - solution that can dissolve more solute at a given temperature
    - ✯ added solute will dissolve
      - ✯ Supersaturated solution - solution that holds more solute than it can dissolve at a given temperature
    - ✯ produced by dissolving solute at a high temperature and allowing it to cool slowly
      - ✯ addition of solute causes precipitation of the excess
- ★ Concentration - the amount of solute compared to solvent
    - ✯ Qualitative descriptions
      - ✯ concentrated - large amount of solute compared to the amount of solvent
        - ✯ example - concentrated orange juice
      - ✯ dilute - small amount of solute compared to the amount of solvent
        - ✯ example - weak coffee

Answer the questions below by circling the number of the correct response

1. A reason why many salts dissociate in water is that water 1. consists of polar molecules, 2. contains ionic bonds, 3. has a linear structure, 4. does not ionize

2. Ammonia gas and hydrogen chloride gas are very soluble in water, which answer best explains the reason for this? 1. water is a good solvent for gases. 2. NH_3, HCl, and H_2O molecules are polar. 3. NH_3, and HCl molecules are very compact. 4. NH_3, HCl, and H_2O molecules are electrically symmetrical.

3. The attraction of water molecules to ions of a solute is 1. hydration, 2. dispersion, 3. ionization, 4. crystallization

4. When an ionic solid dissolves in water, which of the following occurs? 1. ionization of molecules 2. hydration of molecules 3. dissociation of ions 4. formation of ionic bonds with water

5. A reason why many ionic salts dissociate in water is that water
   1. consists of polar molecules 3. contains ionic bonds
   2. has a linear structure 4. does not ionize

6. A solution which contains less solute than should normally dissolve is 1. concentrated, 2. unsaturated, 3. saturated, 4. supersaturated

7. To a solution of NH_4Cl, a crystal of NH_4Cl is added. The crystal falls to the bottom and more solid comes out of the solution. This indicates the original solution was 1. unsaturated, 2. supersaturated, 3. saturated, 4. concentrated

8. A solution in which no more solute can still be added and dissolve is
   1. supersaturated 3. unsaturated
   2. saturated 4. concentrated

9. To a solution of NaCl, a crystal of NaCl is added and the crystal dissolves. The solution must have been 1. supersaturated, 2. saturated, 3. concentrated, 4. unsaturated

10. A solution which contains a maximum amount of solute that can be dissolved under the existing conditions is 1. saturated, 2. unsaturated, 3. dilute, 4. supersaturated

11. Crystals of NaCl, when added to a solution of this salt that is in equilibrium with excess sodium chloride, will 1. dissolve in the solution, 2. cause additional sodium chloride crystals to separate from the solution, 3. form a supersaturated solution, 4. cause no change in the concentration of the solution

12. A saturated solution of which salt would be the most concentrated at 30^oC? (see solubility chart) 1. NaCl, 2. $NaClO_3$, 3. KCl, 4. $KClO_3$

13. Which saturated solution would be most dilute at 0^oC?
    1. KI 3. $NaClO_3$
    2. $NaNO_3$ 4. $KClO_3$

14. Which compound is most soluble in water?(see solubility chart)
    1. silver acetate 3. lead nitrate
    2. silver chloride 4. silver sulfate

15. As the temperature increases from 30^oC to 40^oC, the solubility of potassium nitrate in 100 g of water increases by approximately (see solubility chart)
    1. 5 grams 3. 15 grams
    2. 10 grams 4. 20 grams

16. Which compound is least soluble in 100 grams of water at 10^oC?(see solubility chart)
    1. KNO_3 3. NaCl
    2. KI 4. $KClO_3$

17. A small crystal of the slightly soluble salt calcium sulfate dissolves in a solution of calcium sulfate. The original solution must have been 1. dilute and saturated, 2. concentrated and saturated, 3. dilute and unsaturated, 4. concentrated and unsaturated

18. As the temperature increases and the pressure remains constant, the solubility of a gas in a solution 1. decreases, 2. remains the same, 3. increases, 4. varies directly

19. As the pressure on a gas increases, temperature remaining constant its solubility in water 1. decreases, 2. remains the same, 3. increases, 4. varies inversely

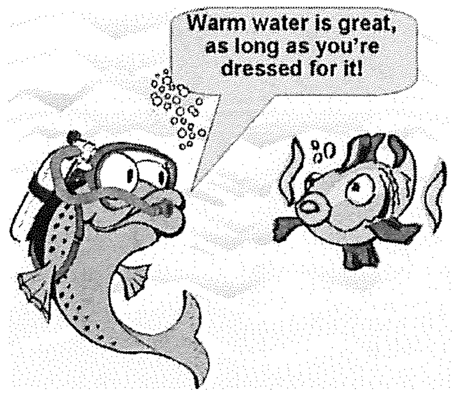
20. Which silver compound is most soluble in water? (see solubility chart) 1. AgCl, 2. AgI, 3. Ag_2SO_4, 4. $AgNO_3$

21. How many grams of KCl are required to saturate 1000 grams of H_2O at 80^oC?(see solubility chart) 1. 390, 2. 500, 2. 800, 4. 1000

Chemistry: Form WS8.1.3A Name ______________________

SOLUTIONS Date ____________ Period ____

Dissolving Solids and Gases



A factory releases clean, warm water into a stream. The stream becomes severely polluted as a result. How does this happen? Fish living in the water depend on dissolved oxygen in order to breathe. Like other gases, oxygen molecules tend to spread out. In order to dissolve them, it is necessary to confine them. Heat speeds the molecules up and makes them spread out more–exactly the opposite of what is needed to dissolve them. As a result, heat drives the oxygen out of the water, causing the fish to die. The dead fish begin to decay. Growing decay bacteria deplete the water of oxygen even further. In this way, clean warm water can pollute a stream. The process of dissolving gases is opposite to the process of dissolving solids because of the differences between gases and solids.

Answer the questions below based on your reading above and on your knowledge of chemistry.

1. A warm can of soda is dropped and bounces down a flight of stairs. When it is opened, carbon dioxide gas coming out of solution causes it to spray all over. Explain the affect of each of the following:

   a. The fact that the soda was warm. ______________________________

   ____

   b. The fact that the soda was dropped and bounced down a flight of stairs. ______________

   ____

   c. The fact that the can was opened. ______________________________

   ____

2. When a gas dissolves, the particles need to be confined. What do the particles of a solid need to do in order to dissolve? ______________________________

3. Sugar is added to a hot cup of coffee and stirred. The sugar dissolves. Explain the affect of each of the following:

   a. The fact that the coffee was hot. ______________________________

   ____

   b. The fact that the coffee was stirred. ______________________________

   ____

Continue ☞

4. Which dissolves faster, a teaspoon of sugar or a sugar cube? Why? ______________________________

____

____

5. A solid is added to water and stirred. Some of it dissolves, but not all. What happens to the rate at which the solid is dissolving between when it was first added and when it stopped dissolving? Explain. (*HINT:* Equilibrium!) ______________________________

____

____

6. The table below lists four factors that may effect the rate at which solids and gases dissolve. Fill in the table by indicating if the rate of dissolving increases, decreases, or is not effected. Then explain why.

Factor	**Affect on Rate of Solution for:**	
	Solid Solutes	**Gaseous Solutes**
Crushing		
Stirring		
Increasing the amount of dissolved solute		
Increasing Temperature		

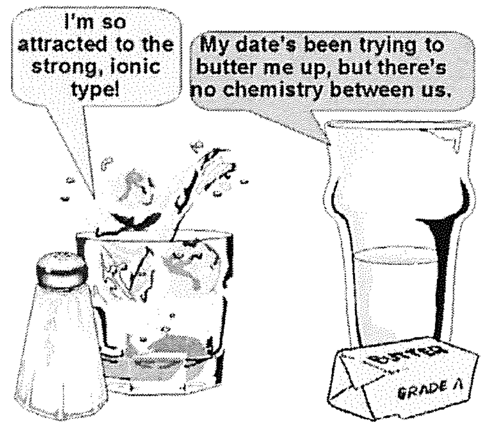
Name ______________________

Date ______________ Period _____

What's in the Water?

A solution is a type of mixture. It consists of two or more kinds of matter each of which retains its own properties. But it is homogeneous. It appears to be only one substance. This is what distinguishes solutions from mechanical mixtures. The substance that is dissolved is the solute. The solvent is the substance that dissolves the solute. It is the continuous phase. For example, salt dissolved in water appears to be a liquid.

Different solutes dissolve best in different solvents. In order for a solvent to dissolve a solute, it must exert forces of attraction on the solute. Polar solvents such as water dissolve polar and ionic solutes well because they exert mutual attractions that cause their particles to intermingle. Of course, not all ionic substances dissolve equally well in water. (See *Table F - Table of Solubilities in Water*) Nonpolar solvents do NOT dissolve polar and ionic substances because there is no attraction between them. For example, oil and water do NOT mix. Nonpolar substances such as fat dissolve in nonpolar solvents such as benzene because the forces of attraction are too weak to prevent the particles from freely intermingling.



Water discusses its tastes.

Answer the questions below based on your reading and on your knowledge of chemistry.

1. Water is mixed with sugar, resulting in a transparent, colorless liquid.
   a. What evidence will there be that this is a mixture rather than a new compound? ______________________
   ____
   b. What evidence shows it is a solution rather than a mechanical mixture? ______________________
   ____
   c. Which is the solute, and which is the solvent? How do you know? ______________________
   ____

2. Why does table salt dissolve in water, while oil and water don't mix? ______________________
   ____
   ____
   ____

3. Based on *Table F* indicate which of the following compounds is water soluble and which is insoluble?

a. Li_2CO_3 ________	e. $(NH_4)_3PO_4$ ________	i. $CuSO_4$ ________
b. $Fe(OH)_3$ ________	f. $Al(ClO_3)_3$ ________	j. KNO_3 ________
c. $CaCrO_4$ ________	g. $PbSO_4$ ________	k. AgCl ________
d. BaS ________	h. NaOH ________	l. $Ba(HCO_3)_2$ ________

NAME:________________________________ DATE:________SECTION:______LAB______

I Have a Solution

Background:

Solutions are homogeneous mixtures. The major component is called **solvent**, and the minor components are called **solute**. If both components in a solution are 50%, the term solute can be assigned to either component. When gas or solid material dissolve in a liquid, the gas or solid material is called the solute. When two liquids dissolve in each other, the major component is called the **solvent** and the minor component is called the **solute**

Some solutions are so common to us that we give them a unique name. A solution of water and sugar is called *syrup*. A solution of carbon dioxide in water is called *seltzer*, and a solution of ammonia gas in water is called *ammonia water*.

A solution is said to be **dilute** if there is less of the solute. The process of adding more solvent to a solution or removing some of the solute is called diluting. A solution is said to be **concentrated** if it has more solute. The process of adding more solute or removing some of the solvent is called concentrating. The **concentration** of a solution is some measurement of how much solute there is in the solution.

Other Types of Mixtures (Not Solutions)

Take a spoonful of dirt and vigorously mix it with a glass of water. As soon as you stop mixing, a portion of the dirt drops to the bottom. Any material that is suspended by the fluid motion alone is only in *temporary suspension*. A portion of the dirt makes a true solution in the water with all of the properties of the below table, but there are some particles, having a diameter roughly between 1 nm and 500 nm, that are suspended in a more lasting fashion. A suspended mixture of particles of this type is called a *colloid*, or *colloidal suspension*, or *colloidal dispersion*.

For colloids or temporary suspensions the phrase *dispersed material* or the word *dispersants* describes the material in suspension, analogous to the solute of a solution. The phrase *dispersing medium* is used for the material of similar function to a solvent in solutions.

Sol is a liquid or solid with a solid dispersed through it.
Foams are liquids or solids with a gas dispersed into them.
Emulsions are liquids or solids with liquids dispersed through them.
Aerosols are colloids with a gas as the dispersing medium and either a solid or liquid dispersant.

PROPERTIES OF SOLUTIONS

1. The particles of solute are the size of individual small molecules or individual small ions. One nanometer is about the maximum diameter for a solute particle.

2. The mixture does not separate on standing. In a gravity environment the solution will not come apart due to any difference in density of the materials in the solution.

3. The mixture does not separate by common fiber filter. The entire solution will pass through the filter.

4. Once it is completely mixed, the mixture is *homogeneous*. If you take a sample of the solution from any point in the solution, the proportions of the materials will be the same.

5. The mixture appears clear rather than cloudy. It may have some color to it, but it seems to be transparent otherwise. The mixture shows no *Tyndall effect*. Light is not scattered by the solution. If you shine a light into the solution, the pathway of the light through the solution is not revealed to an observer out of the pathway.

6. The solute is completely dissolved into the solvent up to a point characteristic of the solvent, solute, and temperature. At a *saturation point* the solvent no longer can dissolve any more of the solute. If there is a saturation point, the point is distinct and characteristic of the type of materials and temperature of the solution.

7. The solution of an ionic material into water will result in an *electrolyte* solution. The ions of solute will separate in water to permit the solution to carry an electric current.

Procedure:

Your group will be give 6 test tubes with possible solutions in them. Based on the introduction and properties of solutions table determine which test tubes contain true solutions.

Test Tube	Mixture	Classification	Reason
A	Air		
B	Brass		
C	Sand & Water		
D	Milk		
E	Oil & Water		
F	Soda		
G	Salt & Water		

Questions:

1) Based on the table above classify smoke in air.

2) Based on the above table classify mayonnaise.

3) Discuss if candle wax will dissolve in oil to form a solution.

4) Which of the following represents a solution?
(1) NaCl (s)
(2) NaCl (l)
(3) NaCl (g)
(4) NaCl (aq)

Reflection:

Describe how a solution is different than just a mixture.

Going Further:

Give an example of a gas/solid solution.

Chemistry: Form WS8.1.2A

SOLUTIONS

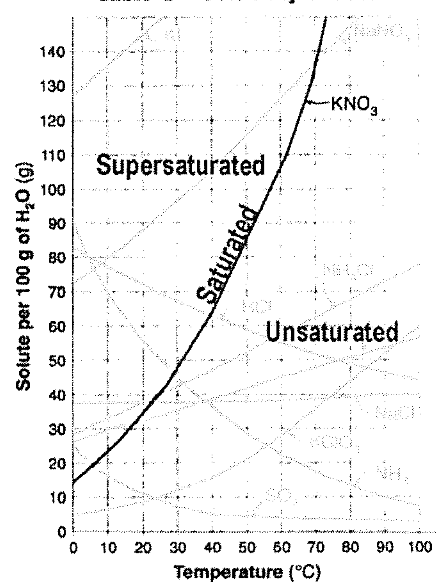
Name ______________________

Date ______________ Period _____

Solubility Curves

The solubility of solid solutes generally increases as temperature increases, while the solubility of gaseous solutes generally decreases as temperature increases. A solution that holds as much solute as can dissolve at a given temperature is saturated. A solution that can dissolve more solute at a given temperature is unsaturated. A solution that holds more solute than can dissolve at a given temperature is supersaturated. The amount of solute that is needed to form a saturated solution at various temperatures can be graphed. This is what is shown in *Table G*. The values in *Table G* are based on solute dissolved in 100 g of water. Since water has a density of 1 g/mL, the graph can be considered to be based on 100 mL of water. A 200 mL sample of water would be able to dissolve twice as much at each temperature.

Table G Solubility Curves



Answer the questions below by referring to *Table G*.

1. The compound which is the most soluble at 20°C is ______________.
2. The compound which is the least soluble at 10°C is ______________.
3. The compound which is the least soluble at 80°C is ______________.
4. The number of grams of potassium nitrate needed to saturate 100 mL of water at 70°C is ______________.
5. The formulas of the compounds which vary inversely with the temperature are ______________, ______________ and ______________.
6. One hundred mL of a sodium nitrate solution is saturated at 10°C. How many additional grams are needed to saturate the solution at 50°C? ______________
7. One hundred mL of a saturate KCl solution at 80°C will precipitate 10 grams of salt when cooled to what temperature? ______________
8. The two salts that have the same degree of solubility at 70°C are ______________ and ______________.
9. The salt with a solubility is least affected by a change in temperature is ______________.
10. The salt that has the greatest increase in solubility in the temperature range between 30°C and 50°C is ______________.
11. The number of grams of sodium nitrate that must be added to <u>50 mL</u> of water to produce a saturated solution at 50°C is ______________.
12. A saturated solution of potassium chlorate is made at 10°C by dissolving the correct mass of salt in 100 mL of water. When the solution is heated to 90°C, how many grams must be added to saturate the solution? ______________

Continue ☞

13. At what temperature do saturated solutions of sodium chloride and potassium chloride contain the same mass of solute per 100 mL of water? ______________________________

14. A saturated solution of potassium nitrate is prepared at 60°C using <u>200 mL</u> of water. If the solution is cooled to 30°C, how many grams will precipitate out of the solution? ______________________________

15. How many more grams of ammonia can be dissolved in 100 mL of water at 10°C than at 90°C? ______________

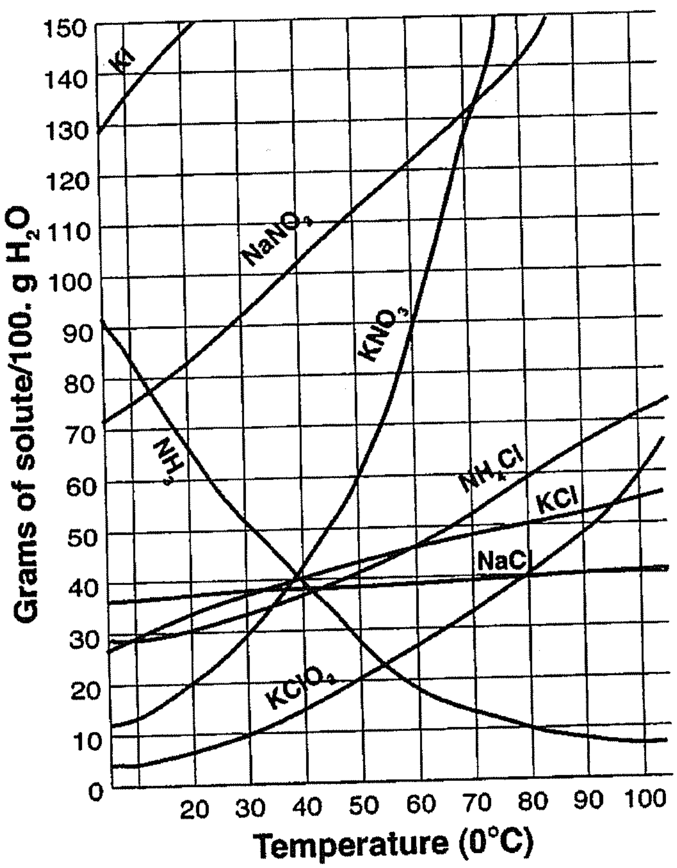
16. A saturated solution of sodium nitrate in 100 mL of water at 40°C is heated to 50°C. The rate of increase in solubility grams per degree is ______________________________.

17. Thirty grams of KCl is dissolved in 100 mL of water at 45°C. The number of additional grams of KCl that would be needed to make the solution saturated at 80°C is ______________________________.

Name ____________________

SOLUBILITY CURVES

Answer the following questions based on the solubility curve below.



1. Which salt is least soluble in water at 20° C? ____________

2. How many grams of potassium chloride can be dissolved in 200 g of water at 80° C?

   ____

3. At 40° C, how much potassium nitrate can be dissolved in 300 g of water? ____________

4. Which salt shows the least change in solubility from 0° – 100° C?

   ____

5. At 30° C, 90 g of sodium nitrate is dissolved in 100 g of water. Is this solution saturated, unsaturated or supersaturated?

   ____

6. A saturated solution of potassium chlorate is formed from one hundred grams of water. If the saturated solution is cooled from 80° C to 50° C, how many grams of precipitate are formed? ____________

7. What compound shows a decrease in solubility from 0° to 100° C? ____________

8. Which salt is most soluble at 10° C? ____________

9. Which salt is least soluble at 50° C? ____________

10. Which salt is least soluble at 90° C? ____________

NAME:________________________________ DATE:______SECTION_____LAB_____

You Have Nice Curves!

Background:

Solubility curves, like the one shown in the CRT, tell us what mass of solute will dissolve in 100 grams of water over a range of temperatures.

You'll notice that for most substances, solubility increases as temperature increases. As previously discussed in solutions involving liquids and solids typically more solute can be dissolved at higher temperatures. The opposite is true of most gasses dissolving in liquids; at higher temperatures less solute can be dissolved.

For all substances as the amount of solvent increases the amount of solute which can be dissolved also increases.

Here's an example of reading the chart.

Find the curve for $KClO_3$.

At 40°C approximately 15 g of $KClO_3$ will dissolve in 100 g of water.

At 40°C approximately 30 g of $KClO_3$ will dissolve in 200 g of water. Two times the water two times the amount of solute.

At 40°C approximately 7.5 g of $KClO_3$ will dissolve in 50 g of water. Halve the water halve the amount of solute.

Problems:

Use Table G in your CRT to answer the following questions

1) List the following salts in order of increasing solubility at 40°C in 100 grams of water: $KClO_3$, KNO_3, $NaNO_3$, KCl, and $NaCl$.

2) You dissolve 70 grams of KNO_3 in 100 grams of water at 60°C. Determine if the solution is saturated. If it is not saturated how many more grams are needed to saturated.

3) You dissolve 73 grams of NH_4Cl in 100 grams of water at 90°C. Determine if the solution is saturated. If it is not saturated how many more grams are needed to saturated.

4) Based on the solubility curves which substances are gases? Explain.

5) Determine the number of grams of $NaNO_3$ needed to saturated 200 grams of water at 40°C.

____________g

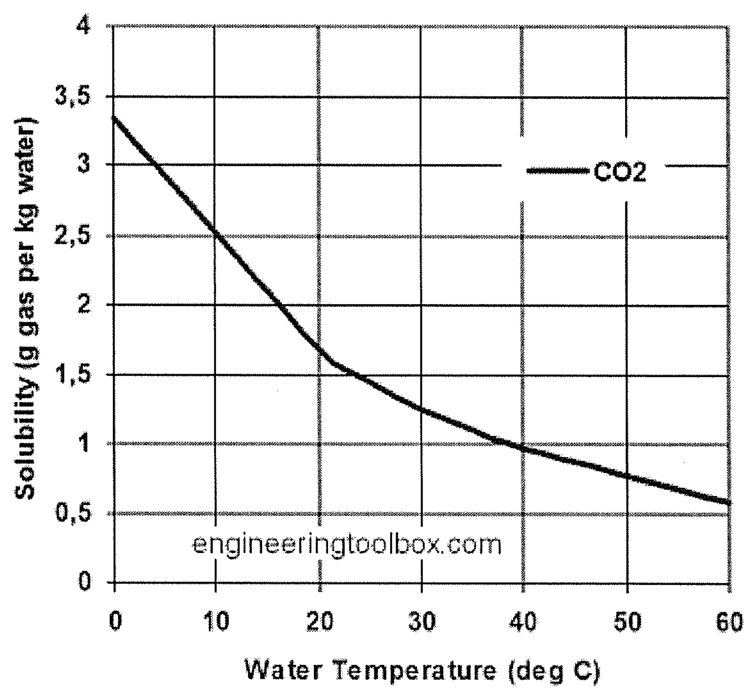
6) Determine how many grams of NH_3 are needed to saturate 50 grams of water at 20°C.

____________g

7) Describe in terms of saturation 65 g of KI dissolved in 50 grams of water at 10°C.
8) Describe in terms of saturation 10 g of SO_2 dissolved in 200 grams of water at 50°C.

9) Describe in terms of saturation 20 g of $KClO_3$ dissolved in 100 grams of water at 40°C.

The chart below shows the solubility curve of carbon dioxide in 1 kilogram of water. That is 1000 grams of water not 100 grams. The commas are supposed to be decimal points.



10) A student dissolves 1.75 grams of CO_2 in 1000 grams of water (1 Kg) at 15°C. In terms of saturation describe this solution.

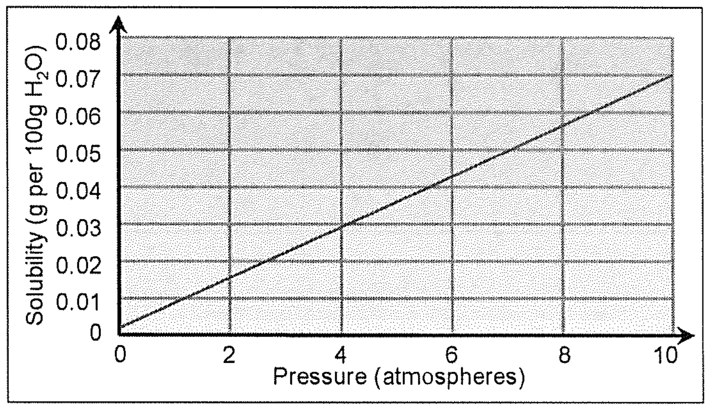
11) Determine the number of grams of CO_2 needed to saturate 1000 grams of water at 35°C.

____________g

12) Describe in terms of polarity why CO_2 gas has such a drastically lower solubility in water and NH_3 gas has a high solubility. You answer must include CO_2, NH_3 and water.

13) Describe what happens to the amount of gas dissolved in water as the temperature increases.

The chart to the right shows the solubility of oxygen gas at various pressures.



14) Describe the effect of increasing pressure on the solubility of a gas.

15) Do you think effect will also affect the solubility of NaCl? Explain.

Reflection:

Describe the method you used to determine the amount of solute needed to saturate a solution with more or less than 100 grams of water used.

Chemistry: Form Ls8.2A — Name ____________

SOLUTIONS — Date ____________ Period ______

Quantities in Solutions

Aim

- to calculate the concentration of a solution, including molarity, percent, and ppm

Notes

Concentration

★ Definition:

$$Concentration = \frac{Mass\ of\ solute(g)}{Volume\ of\ Solvent\ or\ Solution(mL)}$$

★ Molarity

☆ definition: $M = \frac{moles(solute)}{L(solution)}$

☆ related equations

- $M = \frac{g}{GFM \times L}$
- $moles = M \times L$
- $L = \frac{moles}{M}$
- $g = M \times GFM \times L$

☆ examples

Sample Problem 1

Find the molarity of 100 mL of a solution that contains 0.25 moles of dissolved solute.

Step 1: Convert all volumes to liters

$100\ mL \times \frac{1\ L}{1000\ mL} = 0.1\ L$

Step 2: Substitute values into the definitional equation

$M = \frac{moles}{L} = \frac{0.25\ moles}{0.1\ L} = 2.5$ M

Sample Problem 2

Find the molarity of 250 mL of a solution that contains 4 g of dissolved sodium hydroxide (NaOH).

Step 1: Find the GFM

Na	= 23	×	1	= 23
O	= 16	×	1	= 16
H	= 1	×	1	= 1
				40

Step 2: Convert all volumes to liters

$250\ mL \times \frac{1\ L}{1000\ mL} = 0.25\ L$

Step 3: Substitute values into the correct equation

$M = \frac{g}{GFM \times L} = \frac{4\ g}{40\ g/mole \times 0.25\ L} = 0.4\ M$

Sample Problem 3

How many moles of solute are dissolved in 30 mL of a 2 M solution?

Step 1: Convert all volumes to liters

$30\ mL \times \frac{1\ L}{1000\ mL} = 0.03\ L$

Step 2: Substitute values into the correct equation

moles = M × L = (2 moles/L)(0.03 L) = 0.06 moles

Sample Problem 4

How many grams of silver nitrate ($AgNO_3$) are needed to prepare 200 mL of a 0.1 M solution?

Step 1: Find the GFM

Ag	= 108	×	1	= 108
N	= 14	×	1	= 14
O	= 16	×	3	= 48
				170

Step 2: Convert all volumes to liters

$200\ mL \times \frac{1\ L}{1000\ mL} = 0.2\ L$

Step 3: Substitute values into the correct equation

g = M × GFM × L =

(0.1 mole/L)(170g/mole)(0.2 L) = 3.4 g

★ Percent solution and parts per million (ppm)

☆ Percent by mass:

$$percent\ mass = \frac{mass(solute)}{mass(solution)} \times 100\%$$

Sample Problem

What is the percent by mass of a solution containing 2.3 g of ethanol (C_2H_5OH) dissolved in 10.0 g of water?

Step 1: Find the mass of the solution

10.0 g + 2.3 g = 12.3 g

Step 2: Divide the mass of the solute by the mass of the solution and multiply by 100 %

$$percent\ mass = \frac{2.3g}{12.3g} \times 100\% = 19\%$$

☆ Percent by volume:

$$percent\ volume = \frac{volume(solute)}{volume(solution)} \times 100\%$$

Sample Problem

What is the percent by volume of a solution containing 18.2 mL of glycerine ($C_3H_6O_3$) dissolved in 85.0 mL of water?

Step 1: Find the volume of the solution.
18.2 mL + 85.0 mL = 103.2 mL

Step 2: Divide the volume of the solute by the volume of the solution and multiply by 100%

$$percent\ volume = \frac{18.2mL}{103.2mL} \times 100\% = 17.6\%$$

☆ Parts per million

$$ppm = \frac{mass(solute)}{mass(solution)} \times 1{,}000{,}000\,ppm$$

Sample Problem

About 0.0035 g of hydrogen sulfide are dissolved in 10.0 g of water. Express this in parts per million.

Step 1: Find the mass of the solution
10.0 g + 0.0035 g = 10.0035 g

Step 2: Divide the mass of the solute by the mass of the solution and multiply by 1,000,000 ppm.

$$ppm = \frac{0.0035g}{10.0035g} \times 1{,}000{,}000\,ppm = 350\,ppm$$

Answer the questions below by circling the number of the correct response

1. How many grams of H_2SO_4 are contained in 1.00 liter of 0.500 M sulfuric acid? (1) 22.4 (2) 98.0 (3) 49.0 (4) 196
2. In a 2.0 M solution of KOH, how many moles of KOH are contained in 500 milliliters of the solution? (1) 1 (2) 2 (3) 0.5 (4) 4
3. If 0.25 mole of sodium chloride is dissolved in a liter of solution, the molarity of the solution would be (1) 1M (2) 0.50 M (3) 0.25 M (4) 0.125 M
4. If 0.5 liter of water is added to 0.5 liter of 2.0 m KBr solution, the molarity of the resulting solution will be (1) 1.0 (2) 2.0 (3) 0.5 (4) 1.5
5. 29 grams of NaCl are added to enough water to make 1,000. ml of solution. What is the molarity of the solution? (1) 1.00 M (2) 0.29 M (3) 0.50 M (4) 5.00 M
6. What is the molarity of a solution of hydrochloric acid that contains 3.65 grams of HCl dissolved in 1.0 liter of solution? (1) 0.10 M (2) 0.20 M (3) 0.80 M (4) 0.40 M
7. A 1 M solution contains 40 grams of a compound in 500 ml of solution. What is the molecular mass of this compound? (1) 20 (2) 40 (3) 60 (4) 80
8. A 500 ml solution containing 28 grams of KOH is diluted with water to 1,000. ml. What is the molarity of the resulting solution? (1) 1.0 M (2) 2.0 M (3) 0.25 M (4) 0.50 M
9. One liter of a sodium hydroxide solution contains 100 grams of NaOH. The molarity of the solution is (1) 1.0 M (2) 0.25 M (3) 2.5 M (4) 0.50 M
10. When 20.0 grams of NaOH is dissolved in 500 mL of solution, the concentration of the solution is (1) 1.0 M (2) 20 M (3) 0.50 M (4) 4.0 M
11. If 49 grams of pure H_2SO_4 are added to enough water to make 1,000 ml of solution, what is the molarity of the solution? (1) 1.0 M (2) 0.25 M (3) 0.50 M (4) 0.10 M
12. The number of moles of KCl in 1,000 ml of 3 molar solution is (1) 1 (2) 2 (3) 3 (4) 1.5
13. How many moles of H_2SO_4 are present in 250 mL of a 2.00 M solution? (1) 0.50 (2) 2.00 (3) 1.25 (4) 8.00
14. If 500 mL of 1.0 M H_2SO_4 is diluted with H_2O to a new volume of 1,000 mL, the molarity of the new solution is (1) 1.0 (2) 2.0 (3) 0.25 (4) 0.50
15. One liter of a solution of nitric acid contains 126 grams of solution. The molarity of the solution is (1) 1.00 (2) 2.00 (3) 1.26 (4) 0.500
16. How much ethanol C_2H_5OH must be added to water to make 1.0 liter of 0.5 molar solution of ethanol? (1) 0.5 gram (2) 46 grams (3) 23 grams (4) 92 grams
17. What mass of NaOH (formula mass = 40 g.) is needed to prepare 500 mL of 0.50 M solution? (1) 10. grams (2) 20 grams (3) 25 grams (4) 40 grams
18. Two liters of a solution of sulfuric acid contain 98 grams of H_2SO_4. The molarity of this solution is (1) 1.0 (2) 2.0 (2) 0.50 (4) 1.5
19. How many moles of $AgNO_3$ are dissolved in 10 mL of a 1 M $AgNO_3$? (1) 1 (2) 0.1 (3) 0.01 (4) 0.001

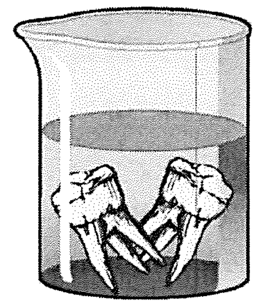
Name ____________________

Date ______________ Period _____

Molarity

One of the most useful measures of concentration in chemistry is molarity (M). Molarity is the number of moles of solute per liter of solution. A two molar (2 M) solution contains two moles of solute per liter of solution.

$$M = \frac{moles(solute)}{L(solution)}$$



A two molar solution

Recall that the number of moles is determined by dividing the number of grams by the gram formula mass (GFM). There are a number of formulas for calculation that come from these relationships:

- $M = \frac{g}{GFM \times L}$
- $moles = M \times L$
- $g = M \times GFM \times L$

Below are some sample problems that show how to apply these formulas.

Sample Problem 1

Find the molarity of 100. mL of a solution that contains 0.25 moles of dissolved solute.

Step 1: Convert all volumes to liters

$$100.mL \times \frac{0.001L}{1mL} = 0.100L$$

Step 2: Substitute values into the definitional equation

$$M = \frac{mol}{L} = \frac{0.25mol}{0.100L} = 2.5M$$

Sample Problem 2

Find the molarity of 500. mL of a solution that contains 4.9 g of dissolved sulfuric acid (H_2SO_4).

Step 1: Find the GFM

H	= 1	× 2	=	2
S	= 32	× 1	=	32
O	= 16	× 4	=	64
				98

Step 2: Convert all volumes to liters

$$500.mL \times \frac{0.001L}{1mL} = 0.500L$$

Step 3: Substitute values into the correct equation

$$M = \frac{g}{GFM \times L} = \frac{4.9g}{(98\,{}^{g}/_{mol})(0.500L)} = 0.10M$$

Sample Problem 3

How many moles of solute are dissolved in 250. mL of a 3.0 M solution?

Step 1: Convert all volumes to liters

$$250.mL \times \frac{0.001L}{1mL} = 0.250L$$

Step 2: Substitute values into the correct equation

$$mol = M \times L = (3.0\,{}^{mol}/_{L})(0.250L) = 0.75mol$$

Sample Problem 4

How many grams of sodium carbonate(Na_2CO_3) are needed to prepare 250 mL of a 0.10 M solution?

Step 1: Find the GFM

Na	= 23	×	2	=	46
C	= 12	×	1	=	12
O	= 16	×	3	=	48
					106

Step 2: Convert all volumes to liters

$$250.mL \times \frac{0.001L}{1mL} = 0.250L$$

Step 3: Substitute values into the correct equation

$$g = M \times L \times GFM = (0.10\,{}^{mol}/_{L})(106\,{}^{g}/_{mol})(0.250L) = 2.7g$$

Answer the questions below based on the reading and the sample problems on the previous page.

1. Determine the molarity of 500. mL of a solution with 0.35 mol of dissolved solute.

2. A 200. mL sample of a solution contains 4.0 g of NaOH. What is its molarity?

3. How many grams of KNO_3 are needed to prepare 25 mL of a 2.0 M solution?

4. How many moles of $MgSO_4$ are contained in 50. mL of a 3.0 M solution?

5. How many grams of $CaCl_2$ are dissolved in 80.0 mL of a 0.75 M solution?

6. What is the molarity of 300 mL of a solution that contains 0.60 mol of dissolved ammonia?

7. What is the molarity of 5.0 L of a solution containing 200. g of dissolved $CaCO_3$?

8. How many grams of NaCl are needed to prepare 500. mL of a 0.400 M solution?

9. How many moles of solute are contained in 3.0 L of a 1.5 M solution?

10. What is the molarity of 750 mL of a solution that contains 40.0 g of dissolved $CuSO_4$?

MOLARITY (M)

Name ______________________

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

Solve the problems below.

1. What is the molarity of a solution in which 58 g of NaCl are dissolved in 1.0 L of solution?

   ____

2. What is the molarity of a solution in which 10.0 g of $AgNO_3$ is dissolved in 500. mL of solution?

   ____

3. How many grams of KNO_3 should be used to prepare 2.00 L of a 0.500 M solution?

   ____

4. To what volume should 5.0 g of KCl be diluted in order to prepare a 0.25 M solution?

   ____

5. How many grams of $CuSO_4 \bullet 5H_2O$ are needed to prepare 100. mL of a 0.10 M solution?

   ____

Name ______________

MOLARITY BY DILUTION

Acids are usually acquired from chemical supply houses in concentrated form. These acids are diluted to the desired concentration by adding water. Since moles of acid before dilution = moles of acid after dilution, and moles of acid = M x V then, $M_1 \times V_1 = M_2 \times V_2$. Solve the following problems.

1. How much concentrated 18 M sulfuric acid is needed to prepare 250 mL of a 6.0 M solution?

   ____

2. How much concentrated 12 M hydrochloric acid is needed to prepare 100 mL of a 2.0 M solution?

   ____

3. To what volume should 25 mL of 15 M nitric acid be diluted to prepare a 3.0 M solution?

   ____

4. To how much water should 50. mL of 12 M hydrochloric acid be added to produce a 4.0 M solution?

   ____

5. To how much water should 100. mL of 18 M sulfuric acid be added to prepare a 1.5 M solution?

   ____

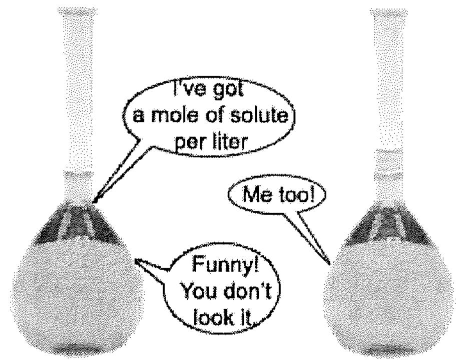
Name ______________________

Date ______________ Period _____

Molality

There are two basic ways to prepare solutions quantitatively. Keep in mind that solute takes up space. Even when the solute is completely dissolved, it affects the volume of the solution. Preparing solutions based on measuring the amount of solute per volume of solution is useful, because it is possible to measure out a sample of the solution and figure out how much solute you have. This is the type of measure that *molarity* is. Understanding how solute affects the solution, however, requires a different type of measure. To compare the solubility of different substances, it is necessary to know how much *solvent* is used compared to solute. To understand how dissolved solute affects the freezing point or boiling point of a liquid, it is necessary to know how much liquid solvent you have for a given amount of solute. *Molality* is very similar to molarity, except that it compares the moles of solute to kilograms of solvent instead of liters of solution. It is abbreviated by lower case *m*.

$$\boxed{m = \frac{mol(solute)}{kg(solvent)}}$$



Competition among molar and molal solutions

Sample Problem 1	Sample Problem 2
Find the molality of a solution that contains 0.45 moles of solute dissolved in 300. g of water. **Step 1**: Convert the amount of solvent to kilograms $300.\,g \times \frac{1kg}{1000g} = 0.300kg$ **Step 2**: Substitute values into the definitional equation $m = \frac{0.45mol}{0.300kg} = 1.5m$	**Find the molality of a solution that contains 25.57 g of sodium chloride dissolved in 250. g of water.** **Step 1**: Find the GFM Na = 22.99 × 1 = 22.99 Cl = 35.45 × 1 = <u>35.45</u> 58.44 **Step 2**: Convert the mass of solute to moles. $(25.57g)\left(\frac{1mol}{58.44g}\right) = 0.4375mol$ **Step 3**: Convert the amount of solvent to kilograms. $250.\,g \times \frac{1kg}{1000g} = 0.250kg$ **Step 4**: Substitute values into the definitional equation. $m = \frac{0.4375mol}{0.250kg} = 1.75m$

Answer the questions below.

1. Find the molality of a solution that contains 225 g of $Ca(NO_3)_2$ dissolved in 400. g of water.
2. Find the molality of a solution that contains 0.663 mol of solute dissolved in 300. g of water.
3. Find the molality of a solution that contains 1.25 kg of KBr dissolved in 2.45 kg of water.

MOLALITY (m)

Name ___________________

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

Solve the problems below.

1. What is the molality of a solution in which 3.0 moles of NaCl is dissolved in 1.5 Kg of water?

   ____

2. What is the molality of a solution in which 25 g of NaCl is dissolved in 2.0 Kg of water?

   ____

3. What is the molality of a solution in which 15 g of I_2 is dissolved in 500. g of alcohol?

   ____

4. How many grams of I_2 should be added to 750 g of CCl_4 to prepare a 0.020 m solution?

   ____

5. How much water should be added to 5.00 g of KCl to prepare a 0.500 m solution?

   ____

☆ Percent by volume:

$$percent\ volume = \frac{volume(solute)}{volume(solution)} \times 100\%$$

Sample Problem

What is the percent by volume of a solution containing 18.2 mL of glycerine ($C_3H_6O_3$) dissolved in 85.0 mL of water?

Step 1: Find the volume of the solution.
18.2 mL + 85.0 mL = 103.2 mL

Step 2: Divide the volume of the solute by the volume of the solution and multiply by 100%

$$percent\ volume = \frac{18.2mL}{103.2mL} \times 100\% = 17.6\%$$

☆ Parts per million

$$ppm = \frac{mass(solute)}{mass(solution)} \times 1{,}000{,}000\,ppm$$

Sample Problem

About 0.0035 g of hydrogen sulfide are dissolved in 10.0 g of water. Express this in parts per million.

Step 1: Find the mass of the solution
10.0 g + 0.0035 g = 10.0035 g

Step 2: Divide the mass of the solute by the mass of the solution and multiply by 1,000,000 ppm.

$$ppm = \frac{0.0035g}{10.0035g} \times 1{,}000{,}000\,ppm = 350\,ppm$$

Answer the questions below by circling the number of the correct response

1. How many grams of H_2SO_4 are contained in 1.00 liter of 0.500 M sulfuric acid? (1) 22.4 (2) 98.0 (3) 49.0 (4) 196
2. In a 2.0 M solution of KOH, how many moles of KOH are contained in 500 milliliters of the solution? (1) 1 (2) 2 (3) 0.5 (4) 4
3. If 0.25 mole of sodium chloride is dissolved in a liter of solution, the molarity of the solution would be (1) 1M (2) 0.50 M (3) 0.25 M (4) 0.125 M
4. If 0.5 liter of water is added to 0.5 liter of 2.0 m KBr solution, the molarity of the resulting solution will be (1) 1.0 (2) 2.0 (3) 0.5 (4) 1.5
5. 29 grams of NaCl are added to enough water to make 1,000. ml of solution. What is the molarity of the solution? (1) 1.00 M (2) 0.29 M (3) 0.50 M (4) 5.00 M
6. What is the molarity of a solution of hydrochloric acid that contains 3.65 grams of HCl dissolved in 1.0 liter of solution? (1) 0.10 M (2) 0.20 M (3) 0.80 M (4) 0.40 M
7. A 1 M solution contains 40 grams of a compound in 500 ml of solution. What is the molecular mass of this compound? (1) 20 (2) 40 (3) 60 (4) 80
8. A 500 ml solution containing 28 grams of KOH is diluted with water to 1,000. ml. What is the molarity of the resulting solution? (1) 1.0 M (2) 2.0 M (3) 0.25 M (4) 0.50 M
9. One liter of a sodium hydroxide solution contains 100 grams of NaOH. The molarity of the solution is (1) 1.0 M (2) 0.25 M (3) 2.5 M (4) 0.50 M
10. When 20.0 grams of NaOH is dissolved in 500 mL of solution, the concentration of the solution is (1) 1.0 M (2) 20 M (3) 0.50 M (4) 4.0 M
11. If 49 grams of pure H_2SO_4 are added to enough water to make 1,000 ml of solution, what is the molarity of the solution? (1) 1.0 M (2) 0.25 M (3) 0.50 M (4) 0.10 M
12. The number of moles of KCl in 1,000 ml of 3 molar solution is (1) 1 (2) 2 (3) 3 (4) 1.5
13. How many moles of H_2SO_4 are present in 250 mL of a 2.00 M solution? (1) 0.50 (2) 2.00 (3) 1.25 (4) 8.00
14. If 500 mL of 1.0 M H_2SO_4 is diluted with H_2O to a new volume of 1,000 mL, the molarity of the new solution is (1) 1.0 (2) 2.0 (3) 0.25 (4) 0.50
15. One liter of a solution of nitric acid contains 126 grams of solution. The molarity of the solution is (1) 1.00 (2) 2.00 (3) 1.26 (4) 0.500
16. How much ethanol C_2H_5OH must be added to water to make 1.0 liter of 0.5 molar solution of ethanol? (1) 0.5 gram (2) 46 grams (3) 23 grams (4) 92 grams
17. What mass of NaOH (formula mass = 40 g.) is needed to prepare 500 mL of 0.50 M solution? (1) 10. grams (2) 20 grams (3) 25 grams (4) 40 grams
18. Two liters of a solution of sulfuric acid contain 98 grams of H_2SO_4. The molarity of this solution is (1) 1.0 (2) 2.0 (2) 0.50 (4) 1.5
19. How many moles of $AgNO_3$ are dissolved in 10 mL of a 1 M $AgNO_3$? (1) 1 (2) 0.1 (3) 0.01 (4) 0.001

ELECTROLYTES

Name ____________________

Electrolytes are substances that break up (dissociate or ionize) in water to produce ions. These ions are capable of conducting an electric current.

Generally, electrolytes consist of acids, bases and salts (ionic compounds). Nonelectrolytes are usually covalent compounds, with the exception of acids.

Classify the following compounds as either an electrolyte or a nonelectrolyte.

Compound	Electrolyte	Nonelectrolyte
1. NaCl		
2. CH_3OH (methyl alcohol)		
3. $C_3H_5(OH)_3$ (glycerol)		
4. HCl		
5. $C_6H_{12}O_6$ (sugar)		
6. NaOH		
7. C_2H_5OH (ethyl alcohol)		
8. CH_3COOH (acetic acid)		
9. NH_4OH ($NH_3 + H_2O$)		
10. H_2SO_4		

SOLUBILITY (POLAR VS. NONPOLAR)

Name ____________________

Generally, "like dissolves like." Polar molecules dissolve other polar molecules and ionic compounds. Nonpolar molecules dissolve other nonpolar molecules. Alcohols, which have characteristics of both, tend to dissolve in both types of solvents, but will not dissolve ionic solids.

Check the appropriate columns as to whether the solute is soluble in a polar or nonpolar solvent.

SOLUTES	SOLVENTS		
	Water	CCl_4	Alcohol
1. NaCl			
2. I_2			
3. ethanol			
4. benzene			
5. Br_2			
6. KNO_3			
7. toluene			
8. $Ca(OH)_2$			

NAME:______________________________ DATE:________SECTION_______LAB_______

We Had a Falling Out

Background:

Precipitation is the formation of a solid in a solution during a chemical reaction. When the reaction occurs in a liquid, the solid formed is called the **precipitate**, and the liquid remaining above the solid is called the **supernate**.

Precipitation Reactions involve the formation of an insoluble product as the driving force of chemical change. To determine the presence of this driving force, predict the products of the reaction, being careful to maintain the oxidation numbers of the cations and anions as you write the formulas of the products. Then check the solubility rules to determine if you formed an insoluble product. If both products are soluble, no precipitation reaction will occur. The ions which are not involved in the precipitation reaction are called **spectator ions**. When writing out a chemical equation it is important to include the states of matter in the equation.

Equations written to represent precipitation reactions can be written in one of three ways:

1) **Molecular Equations**: All reactants and products are written as if they are molecules.

$$AgNO_3\,(aq) + KCl\,(aq) \rightarrow \downarrow AgCl\,(s) + KNO_3\,(aq)$$

2) **Ionic Equations:** All reactants and products that are soluble are written as ions, only the precipitate is written as if it were a molecule.

$$Ag^+\,(aq) \quad NO_3^-\,(aq) + K^+\,(aq) + Cl^-\,(aq) \rightarrow \downarrow AgCl\,(s) + K^+\,(aq) + NO_3^-\,(aq)$$

3) **Net Ionic Equations:** Only the reactants and product taking part in the reaction are written in the equation, the reactants as ions, and the product as a molecule. *Spectator ions* are not included in the equation

$$Ag^+\,(aq) \quad NO_3^-\,(aq) \rightarrow \downarrow AgCl\,(s)$$

Procedure:

1) Write out the ions for each of the reactants given. The start of the ionic equation.
2) Use Table F to determine if any of the ions will combine to make an insoluble product, a precipitate.
3) If a precipitate is formed finish the ionic equation by writing the precipitate as a solid and the spectators as ions. Remember to balance the ionic charges.
4) Complete the molecular equation.
5) Write out the net ionic equation, just showing the ions involved in the precipitation reaction.

Problems:

1) Copper (I) Chloride + Sodium Hydroxide →

Balanced Equation:

$CuCl$ (aq) + $NaOH$ (aq) →

____

Ionic Equation:

____

Net Ionic Equation:

____

______________________ Spectator Ions

______________________ Precipitate

2) Calcium Bromide + Sodium Sulfate →

Balanced Equation:

$CaBr_2$ (aq) + Na_2SO_4 (aq) →

____

Ionic Equation:

____

Net Ionic Equation:

____

______________________ Spectator Ions

______________________ Precipitate

3) Cobalt (II) Chloride + Potassium Carbonate →

Balanced Equation:

$CoCl_2$ (aq) + K_2CO_3 (aq) →

____

Ionic Equation:

____

Net Ionic Equation:

____

____________________ Spectator Ions

____________________ Precipitate

4) Sodium Nitrate + Silver Perchlorate →

Balanced Equation:

$NaNO_3$ (aq) + $AgClO_4$ (aq) →

____

Ionic Equation:

____

Net Ionic Equation:

____

____________________ Spectator Ions

____________________ Precipitate

5) Iron (III) Chloride + Rubidium Hydroxide →

Balanced Equation:

$FeCl_3$ (aq) + RbOH (aq) → ______________________________

Ionic Equation:

____

Net Ionic Equation:

____

______________ Spectator Ions

______________ Precipitate

6) Lithium Phosphate + Magnesium Chloride →

Balanced Equation:

Li_3PO_4 (aq) + $MgCl_2$ (aq) → ______________________________

Ionic Equation:

____

Net Ionic Equation:

____

______________ Spectator Ions

______________ Precipitate

7) Strontium Nitrate + Sodium Sulfate →

Balanced Equation:

____

Ionic Equation:

____

Net Ionic Equation:

____

__________________ Spectator Ions

__________________ Precipitate

8) Lead (II) Acetate + Sodium Chloride →

Balanced Equation:

____

Ionic Equation:

____

Net Ionic Equation:

____

__________________ Spectator Ions

__________________ Precipitate

Use the following information to answer questions 9 through 13.

A saturated solution of potassium nitrate KNO_3 is prepared at 50°C using 100 grams of water. As the solution is cooled to room temperature 20°C a precipitate settles out of the solution. The resulting solution is saturated.

9) Determine which substance in the solution is the solute and solvent.

Solute:______________ Solvent:____________

10) Explain why the KNO_3 is precipitating out of the solution as it cools.

11) Approximately how many grams of KNO3 precipitated out of the original solution? Explain.

____________g

12) Describe a something you could do to make the precipitate re-dissolve in the solution.

13) Describe one process you could use to remove the precipitate from the solution.

14) The following reactions can be classified as

$$AgNO_3\ (aq) + KCl\ (aq) \rightarrow \downarrow AgCl\ (s) + KNO_3\ (aq)$$

(1) Synthesis
(2) Decomposition
(3) Single Replacement
(4) Double Replacement

15) Which of the following compounds will form an insoluble solid when aqueous solutions Sodium Sulfide and Zinc Chloride react?
(1) Na_2S
(2) NaCl
(3) $ZnCl_2$
(4) ZnS

Reflection: Describe how to use Table F to predict the results of a precipitation reaction.

Chemistry: Form N8.3A

SOLUTIONS

Name ______________________

Date ______________ Period ____

Colligative Properties

Aim

- to explain boiling point elevation and freezing point depression

Notes

- Colligative properties - effect of solute on solvent due to the number of particles
  - Nature of colligative properties
    - Not affected by the properties of the solute, but only by the number of particles
    - Electrolytes dissociate producing more particles per mole than nonelectrolytes
      - therefor electrolytes produce larger colligative affects than nonelectrolytes
  - Examples
    - Boiling point elevation - nonvolatile solute reduces the vapor pressure of water, raising the boiling point
      - molal boiling point elevation = 0.512°C/m
    - Freezing point depression - the presence of solute interferes with crystallization, lowering the freezing point
      - molal freezing point depression = 1.86°C/m

Answer the questions below by circling the number of the correct response

1. A pupil dissolved 180.00 grams of $C_6H_{12}O_6$ in 1,000.0 grams of water and then heated the solution until it boiled. What was the boiling point of the $C_6H_{12}O_6$ solution? (air pressure is 1 atmosphere) (1) 98.96°C (2) 100.52°C (3) 99.48°C (4) 101.04°C

2. One mole of an ionic salt will usually depress the freezing point of water to a greater extent than one mole of a soluble organic substance because the ionic salt
   (1) will produce more particles in solution
   (2) is more easily hydrated
   (3) has a higher melting point
   (4) has a higher molecular mass

The solutions described in questions 3 - 9 have the same concentration of dissolved solute.

3. Which solution would have the lowest freezing point? (1) NaCl(aq) (2) HCl(aq) (3) KCl(aq) (4) $CaCl_2$(aq)

4. Which solution will have the lowest freezing point? (1) CH_3COOH (2) $C_6H_{12}O_6$ (3) C_2H_5OH (4) H_2SO_4

5. Which water solution will have the lowest freezing point? (1) $BaCl_2$ (2) NaCl (3) $C_3H_5(OH)_3$ (4) CH_3COOH

6. Which water solution will have the lowest freezing point? (1) $CaCl_2$ (2) $C_{12}H_{22}O_{11}$ (3) NaCl (4) CH_3COOH

7. Which solution will have the highest boiling point?
   (1) KNO_3 (3) $Mg(NO_3)_2$
   (2) $Al(NO_3)_3$ (4) NH_4NO_3

8. Which solution has the lowest freezing point?
   (1) acetic acid (3) nitrous acid
   (2) potassium hydroxide (4) ammonium hydroxide

9. Which water solution will have the highest freezing point?
   (1) $CaCl_2$ (3) $C_{12}H_{22}O_{11}$
   (2) NaCl (4) CH_3COOH

---

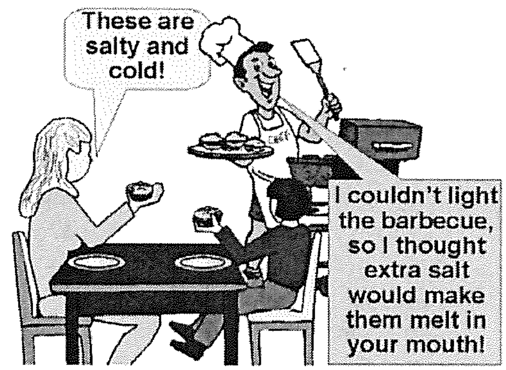
10. If 46.0 grams of ethanol C_2H_5OH are completely dissolved in 1,000. g of water, the freezing point of the solution in Celsius is most nearly (1) 3.72 (2) -1.86 (3) 1.86 (4) -3.72

11. The solution with the lowest freezing point would be produced when 1.0 gram of $C_6H_{12}O_6$ is dissolved in
    (1) 18 grams of H_2O (3) 180 grams of H_2O
    (2) 100 grams of H_2O (4) 1,000 grams of H_2O

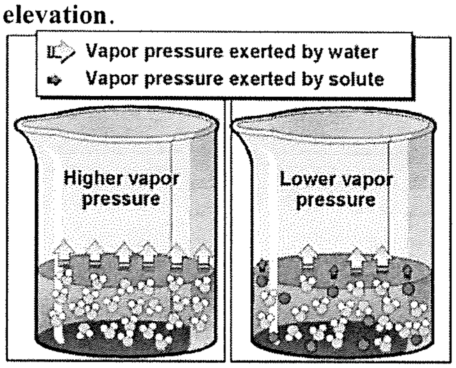
12. What is the total number of grams of $C_6H_{12}O_6$ that must be dissolved in 1,000 grams of water to raise the boiling point 0.52°C? (boiling point elevation constant of H_2O = 0.52°C)
    (1) 9 (3) 18
    (2) 90 (4) 180

Understanding Colligative Properties

After a winter storm, people spread salt on the walks to help melt the ice. Salt reduces the freezing point of water. Actually, any soluble solute reduces the freezing point of water by interfering with crystallization. In this way, antifreeze keeps the water from freezing in an automobile radiator. This phenomenon is called **freezing point depression**. Antifreeze is left in the radiator during the summer. It also prevents the radiator from boiling over by raising the boiling point. Dissolved solute reduces the vapor pressure, raising the boiling point. This is called **boiling point elevation**.



Dad misinterprets freezing point depression.



The amount the freezing point is depressed or the boiling point is raised depends on the concentration of dissolved solute. The higher the concentration of dissolved solute is, the greater the effect on the boiling point or the freezing point is. Only the concentration of the particles of dissolved solute is important. The nature of the solute is not. A mole of dissolved sugar has exactly the same effect on the freezing point and boiling point of 1,000 g of water as a mole of antifreeze because it contains the same number of particles. Ionic compounds dissociate producing more particles per mole. One mole of dissolved sodium chloride, for example, produces one mole of aqueous sodium ions and one mole of aqueous chloride ions for a total of two moles [$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$]. One mole of dissolved sodium chloride, therefore, has twice the effect on the boiling and freezing points of 1,000 g of water as one mole of dissolved sugar. It is not the nature of the solute that matters, but only the concentration of dissolved particles that determines how large the change in freezing point or boiling point will be. Properties of a solution, such as this, which are dependent only on the number of particles in solution, and not on their nature are called **colligative properties.**

Answer the questions below based on your reading and on your knowledge of chemistry.

1. Why are boiling point elevation and freezing point depression considered colligative properties? ______________________

2. Why is salt put on icy roads and sidewalks in the winter? ______________________

3. How will the boiling points of pure water and sea water compare? Why? ______________________

Continue ☞

Solve the following boiling point elevation problems and the freezing point depression problems as shown in the sample problems below. [*NOTE:* At standard pressure, 1 mol of dissolved particles will elevate the boiling point of 1,000 g of water by 0.52 °C and will depress the freezing point of 1,000 g of water by 1.86 °C]

Sample Problem

Find the boiling point of a solution containing 1,000 g of water and 2 mol of dissolved MgF_2.

Step 1: Determine the number of moles of solute particles
$2MgF_2(s) \rightarrow 2Mg^{2+}(aq) + 4F^-(aq)$ $mol = 6$

Step 2: Multiply the boiling point elevation per mole by the number of moles of solute to find the boiling point elevation
BPE = $0.52\,^{°C}/mol \times 3\ mol = 3.12\,°C$

Step 3: Add the boiling point elevation to 100 °C
BP = 100 °C + 3.12 °C = 10 3.12 °C

Sample Problem

Find the freezing point of a solution containing 1,000 g of water and 30 g of dissolved antifreeze ($C_2H_4O_2$).

Step 1: Determine the number of moles of solute particles
C = 12 × 2 = 24
H = 1 × 4 = 4
O = 16 × 2 = 32
60

$$mol = \frac{g}{GFM} = \frac{30g}{60\,^{g}/_{mol}} = 0.5mol$$

Step 2: Multiply the freezing point depression per mole by the number of moles of solute to find the freezing point depression
FPD = $1.86\,^{°C}/mol \times 0.5\ mol = 0.93\,°C$

Step 3: Subtract the freezing point depression from 0 °C
FP = 0 °C – 0.93 °C = – 0.93 °C

4. One mole of dissolved particles elevates the boiling point of 1,000 g of water by 0.52 °C. At standard pressure, what will the boiling point of a solution be if it contains 1,000 g of water and:

   a. 1 mol of antifreeze ($C_2H_4O_2$)? ____________
   b. 1 mol of salt (NaCl)? ____________
   c. 1 mol of ethanol (C_2H_5OH)? ____________
   d. 2 mol of glycerol ($C_3H_6O_3$)? ____________
   e. 2 mol of $CaCl_2(aq)$? ____________
   f. 5 mol of sucrose ($C_{12}H_{22}O_{11}$)? ____________
   g. 1 mol of $KNO_3(aq)$? ____________
   h. 3 mol of $Ba(NO_3)_2(aq)$? ____________
   i. 40 g of $NaOH(aq)$? ____________
   j. 270 g of glucose ($C_6H_{12}O_6$)? ____________

5. One mole of dissolved particles depresses the freezing point of 1,000 g of water by1.86 °C. At standard pressure, what will the freezing point of a solution be if it contains 1,000 g of water and:

   a. 1 mol of glucose ($C_6H_{12}O_6$)? ____________
   b. 1 mol of $BaCl_2(aq)$? ____________
   c. 2 mol of methanol (CH_3OH)? ____________
   d. 3 mol of glycerol ($C_3H_6O_3$)? ____________
   e. 2 mol of $CuSO_4(aq)$? ____________
   f. 4 mol of sucrose ($C_{12}H_{22}O_{11}$)? ____________
   g. 3 mol of $KNO_3(aq)$? ____________
   h. 2 mol of salt (NaCl)? ____________
   i. 150 g of $KHCO_3(aq)$? ____________
   j. 180 g of glucose ($C_6H_{12}O_6$)? ____________

Name ______________

EFFECT OF A SOLUTE ON FREEZING AND BOILING POINTS

We use the following formulas to calculate changes in freezing and boiling point due to the presence of a nonvolatile solute. Freezing point is always lowered, boiling point is always raised.

$\Delta T_F = m \times d.f. \times k_F$ $k_B H_2O = 0.52° C/m$

$\Delta T_B = m \times d.f. \times k_B$ $k_F H_2O = 1.86° C/m$

m = molality of solution

k_F and k_B = constants for particular solvent

d.f. = dissociation factor (how many particles solute breaks up into: for a nonelectrolyte, d.f. = 1)

(Theoretical Dissociation Factor is always greater than observed effect.)

Solve the problems below.

1. What is the new boiling point if 25 g of NaCl is dissolved in 1.0 Kg of water?

____

2. What is the freezing point of the solution in Problem 1?

____

3. What are the new freezing and boiling points of water if 50. g of ethylene glycol (molecular mass = 62 g/mol) is added to 50. g of water?

____

4. When 5.0 g of a nonelectrolyte is added to 25 g of water, the new freezing point is -2.5° C. What is the molecular mass of the unknown compound?

____

NAME:________________________________ DATE:_______SECTION______LAB______

One Goes Down the Other Goes Up

Background:

Colligative properties are those properties of solutions that depend on the number of dissolved particles in solution, but not on the identities of the solutes. For example, the freezing point of salt water is lower than that of pure water, due to the presence of the salt dissolved in the water. To a good approximation, it does not matter whether the salt dissolved in water is sodium chloride or potassium nitrate; if the molar amounts of solute are the same and the number of ions are the same, the freezing points will be the same.

The four commonly studied colligative properties are freezing point depression, boiling point elevation, vapor pressure lowering, and osmotic pressure. Since these properties yield information on the number of solute particles in solution, one can use them to obtain the molecular weight of the solute.

Freezing Point Depression

The presence of a solute lowers the freezing point of a solution relative to that of the pure solvent. For example, pure water freezes at 0°C (32°F); if one dissolves 10 grams of sodium chloride (table salt) in 100 grams of water, the freezing point goes down to −5.9°C (21.4°F). If one uses sucrose (table sugar) instead of sodium chloride, 10 grams in 100 grams of water gives a solution with a freezing point of −0.56°C (31°F). The reason that the salt solution has a lower freezing point than the sugar solution is that there are more particles in 10 grams of sodium chloride than in 10 grams of sucrose. Since sucrose, $C_{12}H_{22}O_{11}$ has a molecular weight of 342.3 grams per mole and sodium chloride has a molecular weight of 58.44 grams per mole, 1 gram of sodium chloride has almost six times as many sodium chloride units as there are sucrose units in a gram of sucrose. In addition, each sodium chloride (an electrolyte) unit comes apart into two ions when dissolved in water. Sucrose is a nonelectrolyte, which means that the solution contains whole $C_{12}H_{22}O_{11}$ molecules. In predicting the expected freezing point of a solution, one must consider not only the number of formula units present, but also the number of ions that result from each formula unit, in the case of ionic compounds. One can calculate the change in freezing point (ΔT_f) relative to the pure solvent using the equation:

$$\Delta T_f = i K_f m$$

where K_f is the freezing point depression constant for the solvent (1.86°C·kg/mol for water), m is the number of moles of solute in solution per kilogram of solvent, and i is the number of ions or molecules present per formula unit (e.g., $i = 2$ for NaCl).

Boiling Point Elevation

The boiling point of a solution is higher than that of the pure solvent. Accordingly, the use of a solution, rather than a pure liquid, in antifreeze serves to keep the mixture from boiling in a hot automobile engine. As with freezing point depression, the effect depends on the number of solute particles present in a given amount of solvent, but not the identity of those particles. If 10 grams of sodium chloride are dissolved in 100 grams of water, the boiling point of the solution is 101.7°C (215.1°F; which is 1.7°C (3.1°F) higher than the boiling point of pure water). The formula used to calculate the change in boiling point (ΔT_b) relative to the pure solvent is similar to that used for freezing point depression:

$$\Delta T_b = i K_b m$$

where K_b is the boiling point elevation constant for the solvent (0.52°C·kg/mol for water), and m and i have the same meanings as in the freezing point depression formula. Note that ΔT_b represents an increase in the boiling point, whereas ΔT_f represents a decrease in the freezing point. As with the freezing point depression formula, this one is most accurate at low solute concentrations.

Questions:

1) Determine the freezing point depression and the actual freezing point of water if you make a 3.0 m KBr aqueous solution.

2) Determine the freezing point depression and the actual boiling point of water if you make a 3.0 m $MgBr_2$ aqueous solution.

3) Describe in terms of number of ions how water's freezing point is affected by the number of ions dissolved in solution.

4) Determine the boiling point elevation and at the actual boiling point of water if you make a 2.0 m ethanol aqueous solution. Remember ethanol is molecular not ionic.

5) Determine the boiling point elevation and at the actual boiling point of water if you make a 4.0 m ethanol aqueous solution. Remember ethanol is molecular not ionic.

6) Describe how water's boiling point is affected by the concentration of solute dissolved in a solution.

7) Determine the freezing point depression and the actual boiling point of water if you dissolve 4.0 moles of NaI in enough water to make 2.0 kg of solution.

8) Determine the boiling point elevation and the actual boiling point of water if you dissolve 5.5 moles of $Al(NO_3)_3$ in enough water to make a 2.3 kg solution.

9) Determine the freezing point depression and the actual freezing point of water if 250. grams of $BeCl_2$ are dissolved in enough water to make 0.100 kg of solution.

10) Determine the boiling point elevation and the actual boiling point of water if 6.5 moles of glycerin are dissolved in enough water to make 8500 grams of solution.

Use the following information to answer questions 11 through 15.

While salt was once a scarce commodity in history, industrialized production has now made salt plentiful. Approximately 51% of world output is now used by cold countries to de-ice roads in winter, both in grit bins and spread by winter service vehicles. Calcium chloride is preferred over sodium chloride, since calcium chloride releases energy upon forming a solution with water; heating any ice or snow it is in contact with. It also lowers the freezing point, depending on the concentration. NaCl does not release heat upon solution; however, it does lower the freezing point. Calcium chloride is thought to be more environmentally friendly than sodium chloride when used to de-ice roads, however a drawback is that it tends to promote corrosion (of vehicles) more so than sodium chloride. NaCl is also more readily available and does not have any special handling or storage requirements, unlike calcium chloride.

11) Determine calcium chloride's chemical formula. ____________

12) In terms of number of ions explain why calcium chloride is a better choice for de-icing roads than NaCl.

13) When calcium chloride dissolves in water the reaction is classified as a
(1) Endothermic
(2) Exothermic
(3) Synthesis
(4) Alpha Decay

14) Which solution has the lowest freezing point? Explain your choice.
(1) 10 g of NaCl in 100 g of water
(2) 20 g of NaCl in 200 g of water
(3) 30 g of NaCl in 100 g of water
(4) 40 g of NaCl in 200 g of water

15) Compared to a 0.1 M aqueous solution of calcium chloride, a 0.8 M aqueous solution of calcium has a
(1) higher boiling point and a higher freezing point
(2) higher boiling point and a lower freezing point
(3) lower boiling point and a higher freezing point
(4) lower boiling point and a lower freezing point

16) Which aqueous solution has the *lowest* freezing point? Explain your choice.
(1) 1.0 M $C_6H_{12}O_6$
(2) 1.0 M C_2H_5OH
(3) 1.0 M $MgCl_2$
(4) 1.0 M NaCl

17) As the ________________dissolved in a solution increases the freezing point decreases
(1) Charge on the ion
(2) Mass of an ion
(3) Amount of ions
(4) Polarity of a molecule

Reflection:

Explain the effect of dissolving a solute on the freezing and boiling point of water.