

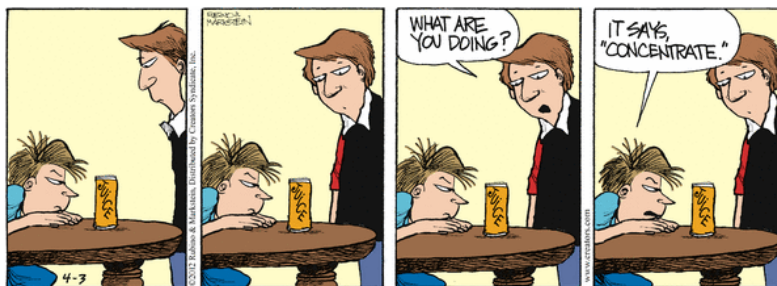
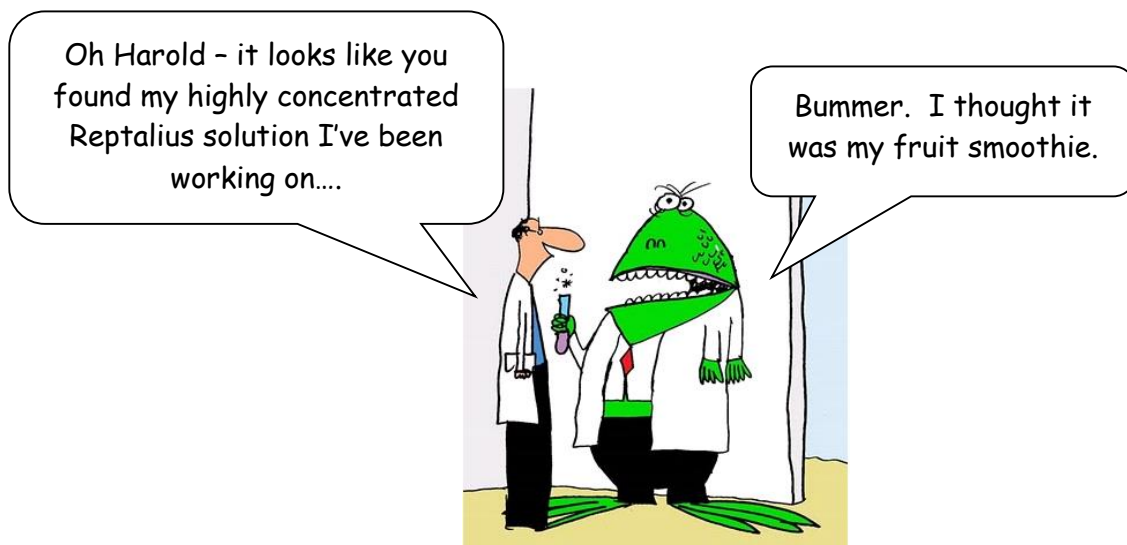
NAME \_\_\_\_\_

PERIOD \_\_\_\_\_

## UNIT 12 NOTES: LIQUIDS & SOLUTIONS

**STUDENT OBJECTIVES:** Your fascinating teachers would like you amazing learners to be able to...

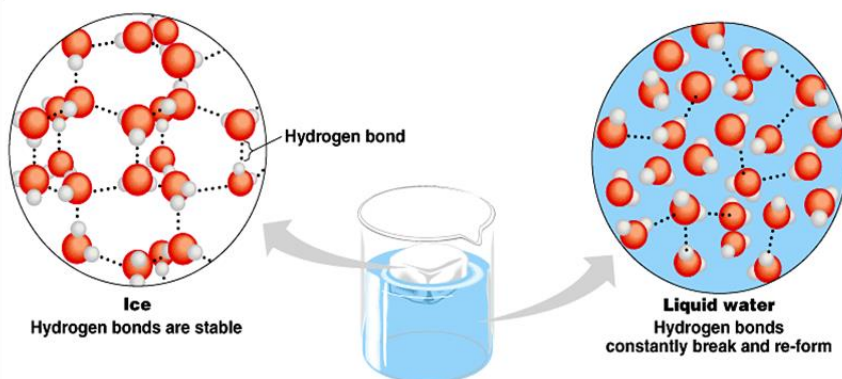
1. Understand properties of liquids, including **cohesion**, **surface tension**, **adhesion**, and **capillary action**.
2. Understand the concepts of **vapor pressure** and **phase equilibrium**.
3. Be able to explain how **boiling point** can be affected by changes in **atmospheric pressure** or by type of **intermolecular forces**.
4. Compare and contrast **solutions**, **colloids**, and **suspensions**.
5. Apply the concepts of molecular polarity to issues of **solubility** and **miscibility**.
6. Develop general rules for solubility through investigations with aqueous solutions.
7. Be able to use solubility rules for a variety of substances.
8. Write the balanced chemical reaction to represent **dissociation**.
9. Determine the number of particles,  $i$ , formed when a substance dissolves in water.
10. Predict whether or not a substance would be an **electrolyte** when dissolved in water.
11. Explain the factors that affect the **rate of solubility** and **amount of solubility** of a solute.
12. Interpret and calculate with **solubility curves**.
13. Compare **unsaturated**, **saturated**, and **supersaturated** solutions.
14. Perform calculations involving **molarity** – both solving for  $M$ , and using within stoichiometry.
15. Evaluate **colligative** properties - **boiling point elevation** and **freezing point depression**.
16. Understand the chemical and biological importance of water and its special structure.



## I. LIQUID PROPERTIES

A. As a reminder, here are some of the qualities that we see in liquids...

- Liquids do sense \_\_\_\_\_ (but not quite as much as solids).
- Liquid molecules are very \_\_\_\_\_ together (way more than gases, but not quite as much as solids).
- Liquid molecules are \_\_\_\_\_ able to be \_\_\_\_\_ or \_\_\_\_\_ to any measureable amount.
- Liquids are \_\_\_\_\_, which allows their molecules to diffuse easily with each other.
- Liquids take the \_\_\_\_\_ of their \_\_\_\_\_, but \_\_\_\_\_ to fill it (meaning, it has a definite volume).
- Liquids are \_\_\_\_\_ than their solid counterparts, with one key, very important exception – \_\_\_\_\_!



The phenomenon with water occurs because of the open shape of the solid crystal structure that forms when Hydrogen Bonding IMFs are stable, as shown in the picture to the left.

**EXAMPLE 12-1.** What are some biological scenarios that are affected by the fact that solid water (ice) is less dense than liquid water?

However, there are many other properties of liquids that we have not discussed!

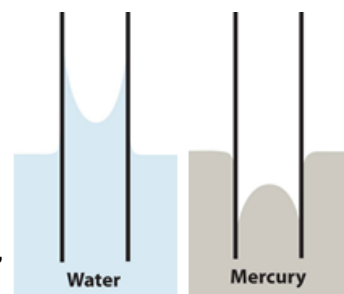
- B. \_\_\_\_\_ – liquids that have \_\_\_\_\_ intermolecular forces tend to stick together. Have you ever noticed that water droplets tend to “bead” together? This is due to water’s strong Hydrogen Bonding IMFs and polarity – the water molecules want to be together instead of apart. If the IMFs are weak, then cohesion will not be observed.

Cohesion also leads to high \_\_\_\_\_ – where the surface of a liquid with high cohesion/IMFs is more difficult to penetrate. Surface tension allows some insects to walk across the surface of water, as the insect’s mass is not great enough to break through the strength of the hydrogen bonding.

- C. \_\_\_\_\_ – this is the \_\_\_\_\_ of a liquid to another substance. When some liquids come in contact with other substances that attract them, the liquid will “rise up” along the surface of that substance. Again, a substance likely to show this phenomenon is water! We see this in action when water makes a “concave meniscus” in a glass tube – the adhesion forces are stronger than the cohesive forces. But, if the cohesive forces are stronger than the adhesion forces, you will get a “convex meniscus”, as seen on the surface of liquid mercury!

The combination of adhesion and cohesion can lead to \_\_\_\_\_. As water attracts and rises up a surface due to adhesion, the molecules will pull along other molecules due to cohesion. This helps water to rise up against gravity. Again, this is due to water’s strong IMFs and polarity.

**EXAMPLE 12-2.** When around us do we see water “rise up” against gravity (capillary action)?



Leggett PreAP Chem Solutions 3 (10:36)

<http://youtu.be/zboci-ck9yQ>

<https://vimeo.com/61856960>

## II. VAPORIZATION OF LIQUIDS

- A. Some review about vaporization...

- \_\_\_\_\_ is the process of vaporization at the boiling point.
- \_\_\_\_\_ is the process of vaporization below the boiling point.
- Vaporization is an \_\_\_\_\_ process, as it requires an input of heat/energy to weaken IMFs to change from a liquid into a gas.

- B. Vaporization takes place when the pressure of the molecules escaping off the surface of the liquid ( $P_{\text{vap}}$ ) is able to overcome the pressure of the air ( $P_{\text{air}}$ ).

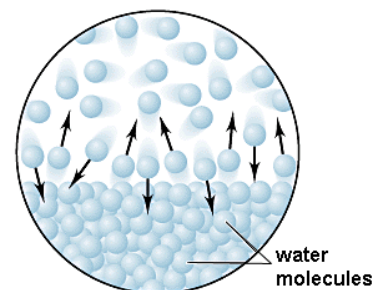
Basically, the liquid molecules need to have enough energy to \_\_\_\_\_ their IMFs and bust through the pressure of the air pressing on the liquid's surface!

### C. VAPOR PRESSURE

At any temperature, all liquids have some \_\_\_\_\_ taking place at the surface. This is because there will always be some molecules that have enough energy to overcome IMFs. Remember that even when you record a temperature, you are recording an average kinetic energy – meaning that some molecules have more energy, and some have less.

What this means is that all liquids have a \_\_\_\_\_ – a pressure exerted by the resulting vapor (gas) of molecules escaping from the surface of the liquid at a particular temperature ( $P_{\text{vap}}$ ).

As temperature increases, the vapor pressure of a liquid will \_\_\_\_\_ due to more molecules having the energy to escape the surface of the liquid! (You saw this trend on the water vapor pressure table we used last unit!)



Note: As temperature increases, more water would be escaping the surface, so vapor pressure would increase.

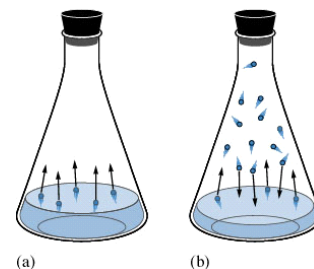
### D. EVAPORATION RATES

If left alone, most of us know that a liquid will eventually vaporize on its own (\_\_\_\_\_). But why do some liquids evaporate more quickly than others?

- Water (relatively speaking) evaporates \_\_\_\_\_. This is because it has \_\_\_\_\_ Hydrogen Bonding IMFs that need a lot of energy to weaken.
- Acetone (nail polish remover) evaporates very \_\_\_\_\_. This is because it mostly only has \_\_\_\_\_ London-Dispersion IMFs that do not need as much energy to break apart into gases – which leads to \_\_\_\_\_ boiling points.
- Liquids that evaporate very easily and quickly due to weak IMFs are said to be \_\_\_\_\_.

### E. VAPORIZATION IN A CLOSED CONTAINER

- In a closed container at a constant temperature, vaporization will happen, but eventually condensation will begin, until both ultimately reach a steady state.
- This means that a \_\_\_\_\_ has been reached – where the \_\_\_\_\_ of vaporization and condensation will be equal.
- This is why when you drink part of a bottle of water, put the cap back on it, and let the bottle sit, that little water droplets form on the upper part of the bottle. The water does not care if it's condensing on the surface of the water or not – it just wants to happen at an equal rate to the vaporization occurring.

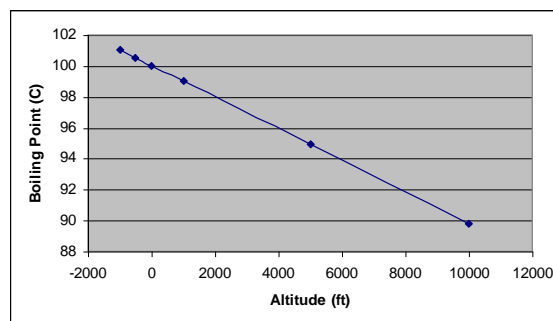


Leggett PreAP Chem Solutions 4 (4:23)

<http://youtu.be/aVJ0q7Zc00Y><https://vimeo.com/61856240>

## F. BOILING POINT AND ELEVATION

- The \_\_\_\_\_ is the boiling temperature when the atmospheric pressure is \_\_\_\_\_. However, the atmospheric pressure is only 1 atm at \_\_\_\_\_! So what happens when the atmospheric pressure is different?
- When you are at higher altitudes, the atmospheric pressure ( $P_{\text{air}}$ ) is \_\_\_\_\_ than 1 atm. Therefore, it does not require as much \_\_\_\_\_ to get the  $P_{\text{vap}}$  to be greater than the  $P_{\text{air}}$ ! So, the boiling point will be \_\_\_\_\_ than normal. The opposite is true at low elevations.
- For example, the normal boiling point of water is \_\_\_\_\_. But, on Mt. Everest, the boiling point is approximately  $70^{\circ}\text{C}$ !
- NOTE:** Elevation only affects the boiling point – it does not affect the freezing/melting point of a substance, as no gas is involved in that phase change. **Remember, only gases are greatly impacted by changes in pressure.**



**EXAMPLE 12-3.** Would food cook faster in an oven at a higher elevation, or at a lower elevation, if the oven is set to the same temperature? Explain.



Dead Sea: -1312 ft

As Altitude  $\uparrow$   
 $P_{\text{atm}} \downarrow$   
 $P_{\text{vap}} = P_{\text{atm}}$  sooner  
 $T_{\text{b}} \downarrow$



Mt Everest: 29,035 ft

Leggett PreAP Chem Solutions 5 (15:22)

<http://youtu.be/3oFbw3Mck1Q><https://vimeo.com/61856241>

## III. SOLUTION VOCABULARY

Before we talk about solution formation, we need to recall some basics about the differences between heterogeneous mixtures and homogeneous mixtures (solutions).

**MIXTURES COMPARISON CHART**

	SOLUTIONS	COLLOIDS	SUSPENSIONS
<b>Type of Mixture</b>	Homogeneous	Heterogeneous	Heterogeneous
<b>Composition</b>	Uniform throughout	Not uniform	Not uniform
<b>Particle Size and Type</b>	Small (.01 – 1 nm), can be atoms, ions, or molecules	Medium (1-1000 nm), dispersed throughout, can be large molecules	Large (over 1000 nm), large particles
<b>Separate on Standing?</b>	No	No	Yes – Particles settle out
<b>Separate by Filtration?</b>	No	No	Yes
<b>Scatter Light?</b>	No	Yes	Yes
<b>Example</b>	Salt Water	Milk	Salad Dressing

In this unit, we will be getting WAY more in-depth with HOMOGENEOUS MIXTURES – also known as solutions. It is important to remember that solutions only contain very small particles dissolved in another substance.

\_\_\_\_\_ - capable of being dissolved.

Dissolving is considered a \_\_\_\_\_ change, since the chemical formulas of the substances do not change.

\_\_\_\_\_ - the dissolving medium in a solution (what is doing the dissolving). *If water is involved, it will be the solvent*; also, for solids into liquids, the liquid is typically the solvent.

\_\_\_\_\_ - the substance dissolved in a solution (what is being dissolved).

**EXAMPLE 12-4.** So, in salt water, what is the solute and what is the solvent?

**Question: Do solutions always have to be a solid dissolved in a liquid?**

NO! We can have many types of solutions, involving all three phases!

(1) Gas / Liquid:

(2) Gas / Gas:

(3) Liquid / Liquid:

(4) Solid / Solid:

**Question: Does everything always dissolve in something else?**

NO! Not necessarily. It depends of if the substances are soluble or not!

\_\_\_\_\_ - when a substance does not dissolve in a solvent

**For liquids, we have special names that refer to their solubility and insolubility.**

\_\_\_\_\_ - liquids that are soluble with each other. (EX: vinegar and water)

\_\_\_\_\_ - liquids that are insoluble with each other (EX: oil and water)

**EXAMPLE 12-5.** Immiscible liquids will not mix, but rather separate into layers. What determines which liquid will be on top of another liquid?

## IV. SOLUBILITY RULES – What will form a solution, and what will not?

A. For molecular (COVALENT) compounds, the rule is \_\_\_\_\_ - that is, particles with similar molecular polarities will dissolve each other.

**Polar solvents will dissolve polar solutes, and non-polar solvents will dissolve non-polar solutes.**

**EXAMPLE 12-6.** Indicate whether the following dissolutions will occur. *DRAWING LEWIS DOT STRUCTURES MAY HELP YOU...*

a. Water with Phosphorus Tribromide?

b. Ammonia (NH<sub>3</sub>) with Methane?

c. Hydrogen gas with Chlorine gas?

d. Water with Carbon Tetrafluoride?

e. Ethane with Propane?

Leggett PreAP Chem Solutions 6 (15:42)

<http://youtu.be/FcBgrVIGKvg>

<https://vimeo.com/61856243>

B. For ionic compounds, the rules aren't quite so clear. The ability to dissolve depends on the relative strength of the solute-solvent interactions.

**FIRST - Ionic compounds will only dissolve only in \_\_\_\_\_ SOLVENTS...they do not dissolve in non-polar substances.**

Even still, there are certain ionic compounds that will dissolve in polar substances, and some that will not. There are special rules for ionic substances in polar solvents (like water)...

### Solubility of Common Ionic Compounds in Water

SOLUBLE COMPOUNDS	EXCEPTIONS
All Group 1 salts	None
All ammonium (NH <sub>4</sub> <sup>+1</sup> ) salts	None
All NO <sub>3</sub> <sup>-1</sup> , ClO <sub>3</sub> <sup>-1</sup> , ClO <sub>4</sub> <sup>-1</sup> , and C <sub>2</sub> H <sub>3</sub> O <sub>2</sub> <sup>-1</sup> salts	None
All Cl <sup>-1</sup> , Br <sup>-1</sup> , I <sup>-1</sup> salts	Ag <sup>+1</sup> , Hg <sub>2</sub> <sup>+2</sup> (mercury (I)), Pb <sup>+2</sup>
All F <sup>-1</sup> salts	Mg <sup>+2</sup> , Ca <sup>+2</sup> , Sr <sup>+2</sup> , Ba <sup>+2</sup> and Pb <sup>+2</sup>
All salts of SO <sub>4</sub> <sup>-2</sup>	Ca <sup>+2</sup> , Sr <sup>+2</sup> , Ba <sup>+2</sup> , Pb <sup>+2</sup> , Ag <sup>+1</sup> , Hg <sub>2</sub> <sup>+2</sup>

INSOLUBLE COMPOUNDS	EXCEPTIONS
All salts of OH <sup>-1</sup>	Group I, NH <sub>4</sub> <sup>+1</sup> , Ba <sup>+2</sup> , Sr <sup>+2</sup> , Ca <sup>+2</sup>
All other monatomic and polyatomic anions	Group I and NH <sub>4</sub> <sup>+1</sup>

**ALSO:** If a compound does not apply to any of the above rules (such as, it contains a polyatomic ion not mentioned above), it is **INSOLUBLE**, unless with an alkali metal cation or ammonium.



Solubility rules are really great! They help us predict precipitates!



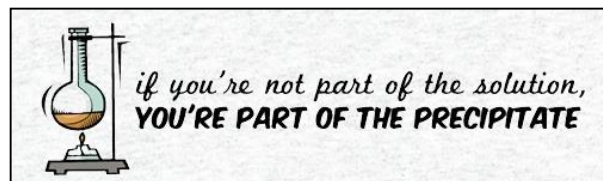
**A SPECIAL NOTE ABOUT WATER:** Because so many substances are either polar covalent or an ionic compound soluble in water, water (being itself polar) will dissolve most substances.

Therefore, we call water the \_\_\_\_\_.

**EXAMPLE 12-7.** For each of the following, use solubility rules to determine if it is soluble in polar solvents or not. Put the correct state symbol next to the compound.

- |                             |                             |
|-----------------------------|-----------------------------|
| a. NaCl                     | b. AgBr                     |
| c. $\text{NH}_4\text{NO}_3$ | d. $\text{CaF}_2$           |
| e. MgO                      | f. $\text{Ca}(\text{OH})_2$ |
| g. $\text{MgCO}_3$          | h. $\text{CaSO}_4$          |

As a reminder, when you have an insoluble solid that collects at the bottom of a solution, we call that solid a \_\_\_\_\_.



### USING SOLUBILITY RULES

If a compound is soluble, it will get the \_\_\_\_\_ state designation, as it will dissolve in water.

If a compound is insoluble, that means that it will not dissolve in water, and will form a \_\_\_\_\_.

The substance will get the \_\_\_\_\_ state designation.

### c. So, to summarize solubility rules...

**Covalent + Covalent:**

**Covalent + Ionic...**

**Non-Polar + Ionic:**

**Polar + Ionic:**

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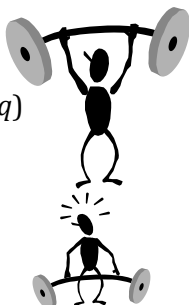
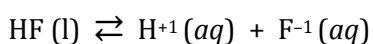
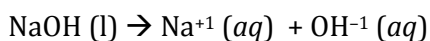
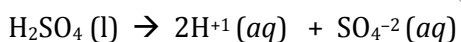
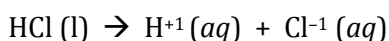
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## V. DISSOCIATION OF SOLUBLE SUBSTANCES

- A. \_\_\_\_\_ is the process of a compound splitting into its ions when dissolved in a polar solvent. Acids, Bases, and Ionic Compounds can all undergo dissociation.

While polar covalent compounds may be soluble in water, they will \_\_\_\_\_ because they are not composed of ions. (They can dissolve – but they do not dissociate into ions.)

- B. For acids & bases, \_\_\_\_\_ acids and bases completely dissociate – it completely dissociates into ions, so we show the dissociation with a one-way arrow. \_\_\_\_\_ acids and bases only **partially** dissociate, so we show them with a double arrow.



**How do I know what's a strong acid or a strong base???**

Strong Acids:

HCl, HBr, HI,  $\text{HNO}_3$ ,  $\text{H}_2\text{SO}_4$ ,  $\text{HClO}_3$ , and  $\text{HClO}_4$

Strong Bases:

$\text{OH}^-$  with a metal from Group 1 or Group 2, except Be & Mg

Every other acid and base is considered to be weak.



**EXAMPLE 12-8.** Write ionization equations for Nitric Acid, Sulfurous Acid, and Strontium Hydroxide dissolving in water.

Nitric Acid:

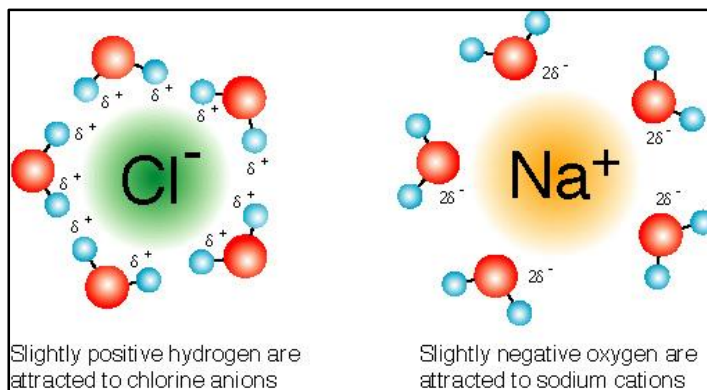
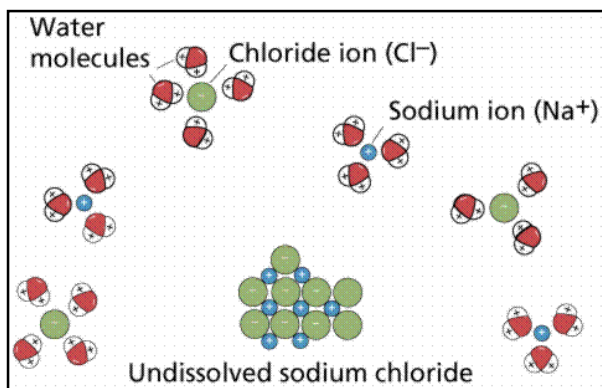
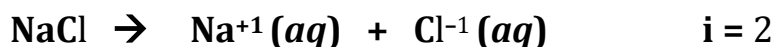
Sulfurous Acid:

Strontium Hydroxide:

C. Soluble ionic compounds will completely dissociate. This is where your solubility rules will come in handy!

**We will use the symbol “i” (lower case letter) to indicate the number of particles formed when a substance dissolves.** For polar molecular substances which dissolve in water, i is always equal to one since they DO NOT dissociate when they dissolve!

EX: sodium chloride is dissolved in water



**EXAMPLE 12-9.** Fill in the information on the following chart.

SUBSTANCE	IONIC OR COVALENT?	SOLUBLE IN WATER?	DISSOCIATION EQUATION	i
Lithium phosphate				
Calcium sulfate				
Aluminum perchlorate				
Copper (II) iodide				
Ammonia (NH <sub>3</sub> )				
Methane				
Ammonium chloride				

Oh, I still hate those polyatomic ions...



## VI. ELECTROLYTES – When will solutions conduct electricity?

An electrolyte is a substance that has been dissolved in water and has dissociated to give a solution of ions... which has the ability to \_\_\_\_\_.

Electrical conduction requires a pathway for the \_\_\_\_\_ of \_\_\_\_\_.... meaning we need charged particles that can move. There are several things that conduct electricity, but to be considered an electrolyte, the substance must contain ions.

- A. \_\_\_\_\_ compounds are \_\_\_\_\_ electrolytes when dissociated... they conduct electricity very well.
- B. Strong acids and bases are \_\_\_\_\_ electrolytes... they conduct electricity very well.
- C. Weak acids and bases are \_\_\_\_\_ electrolytes... they conduct electricity, but poorly.
- D. \_\_\_\_\_ compounds and \_\_\_\_\_ conduct electricity – but they are not dissolved! So they are usually not considered to be electrolytes.
- E. \_\_\_\_\_ are NEVER electrolytes – while some may dissolve, they do not form ions. They are called non-electrolytes.
- F. \_\_\_\_\_ compounds are NEVER electrolytes – they do not dissolve, so they do not form ions. They are also non-electrolytes.

Many people know me as containing "electrolytes"... that just means I contain dissolved, dissociated ions!



**EXAMPLE 12-10.** Tap water will conduct electricity, while distilled water does not. Propose an explanation for this difference.

**EXAMPLE 12-11.** Fill in the information on the following chart.

Substance	Formula?	Ionic or Acid/Base or Covalent?	Soluble?	Strong or Weak or Non- Electrolyte?
Ammonium carbonate				
Copper (II) sulfite				
Barium sulfate				
Magnesium nitrate				
Calcium chloride				

Substance	Formula?	Ionic or Acid/Base or Covalent?	Soluble?	Strong or Weak or Non- Electrolyte?
Hydrochloric Acid				
Acetic Acid				
Methane				
Ammonia				

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## VII. SOLUBILITY – How can we change how much and how quickly we can dissolve?

\_\_\_\_\_ refers to the maximum amount of solute that will dissolve in a specific amount of solvent at a specific temperature and pressure.

### A. Factors that determine the RATE of dissolution (for solids dissolving in liquids):

- \_\_\_\_\_ - by stirring or shaking a substance, you will get more collisions between solute and solvent! (ex: stirring the sugar in iced tea)
- \_\_\_\_\_ - by exposing more of the solute to the solvent, the two will have more collisions! (ex: granulated sugar vs. cubed sugar)
- \_\_\_\_\_ - higher temperature means more kinetic energy, which means more movement... which leads to more collisions!

### B. Factors that affect AMOUNT of solubility (for either gases or solids dissolving in liquids):

- \_\_\_\_\_ - Obviously, the more solvent you have, the more total amount (not percentage) you will be able to dissolve. All tables and figures showing solubility MUST indicate the amount of solvent involved.
- \_\_\_\_\_ - different for solids and gases dissolving in a liquid!!!
  - Solid solutes in water: If you **increase** temperature, the solubility usually \_\_\_\_\_. (ex: Making sweet tea! By heating up the tea, more sugar dissolves.)
  - Gas solutes in water: If you **increase** temperature, the solubility \_\_\_\_\_. (ex: Soda... it has much more carbonation when it is cold than when it gets hot.)
- \_\_\_\_\_ - mainly only affects the solubility of \_\_\_\_\_ in liquids.
  - Gas solutes in water: If you **increase** pressure, the solubility \_\_\_\_\_. (ex: Soda... it has more carbonation when the can or bottle is closed versus when it is open!)

C. Describing the saturation of your solution:

\_\_\_\_\_ - contains the **maximum** amount of dissolved solute for a given amount of solvent at a specific temperature and pressure (you have dissolved everything that a solubility chart says you can).

\_\_\_\_\_ - contains **less** than maximum of dissolved solute for a given temperature and pressure. You can still dissolve more if you want to! If you put more solute into a solution and it completely dissolves, your solution was unsaturated.

\_\_\_\_\_ - Contains **more** dissolved solute than a saturated solution at the same temperature. You have forced more solute to get dissolved! This happens by heating up the substance and cooling it off very slowly. An example of this is Sweet Tea.



Supersaturated solutions contain more solute than normal for that temperature and so are unstable. This unstable state can be disrupted by adding more solute into the solvent. When you do so, you will decrease your solubility down to normal saturated conditions and all of the extra solute that was being dissolved will precipitate (fall) to the bottom of your container. For example, adding more sugar to sweet tea can actually cause your tea to become less sweet because all of the extra sugar that was being dissolved will be disrupted and fall to the bottom of your glass.

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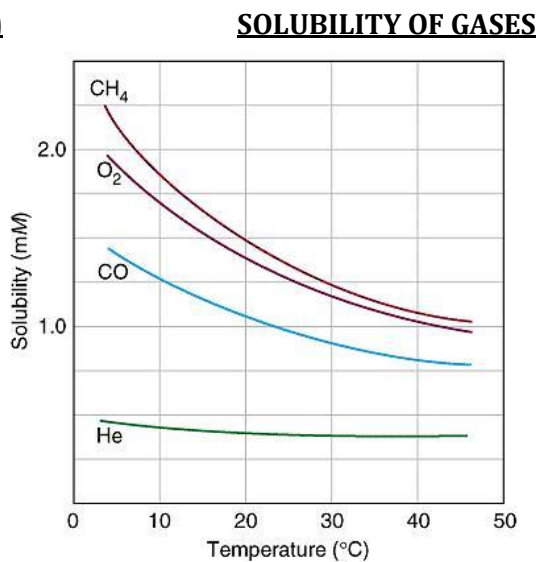
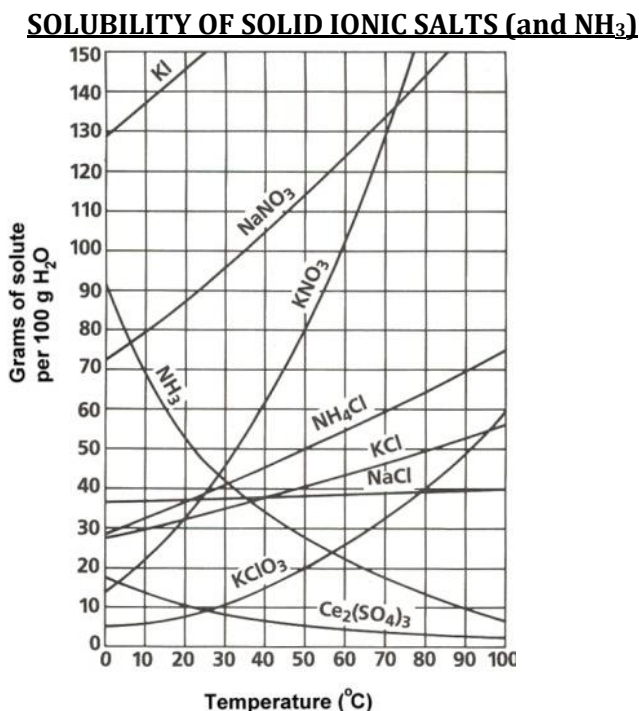
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**EXAMPLE 12-12.** 4 solutions were made using 100.0 mL of water and varying dissolved amounts of NaCl. An additional 30.0 g of NaCl was then added to the solutions, the solutions were stirred for 5 minutes, and the results were recorded. Classify the original solutions as saturated, unsaturated, or supersaturated.

Solutions containing 100.0 mL of water	Amount of NaCl added to solution	Results after stirring for 5 minutes	Original solution would be classified as...
Solution 1	30.0 g NaCl added	0.0 g NaCl at bottom of container	
Solution 2	30.0 g NaCl added	45.0 g NaCl at bottom of container	
Solution 3	30.0 g NaCl added	30.0 g NaCl at bottom of container	
Solution 4	30.0 g NaCl added	10.0 g NaCl at bottom of container	

## VIII. SOLUBILITY CURVES

Below are solubility curves showing the solubility of several different compounds at different temperatures. The solubility chart with the solid compounds is expressed as grams of solute dissolved in 100 grams of water. This can be different depending on your chart. Make sure to be aware of what your chart shows.



**EXAMPLE 12-13.** What do you notice in general about the solubility of solids with increasing temperature?

**EXAMPLE 12-14.** What do you notice in general about the solubility of gases with increasing temperature?

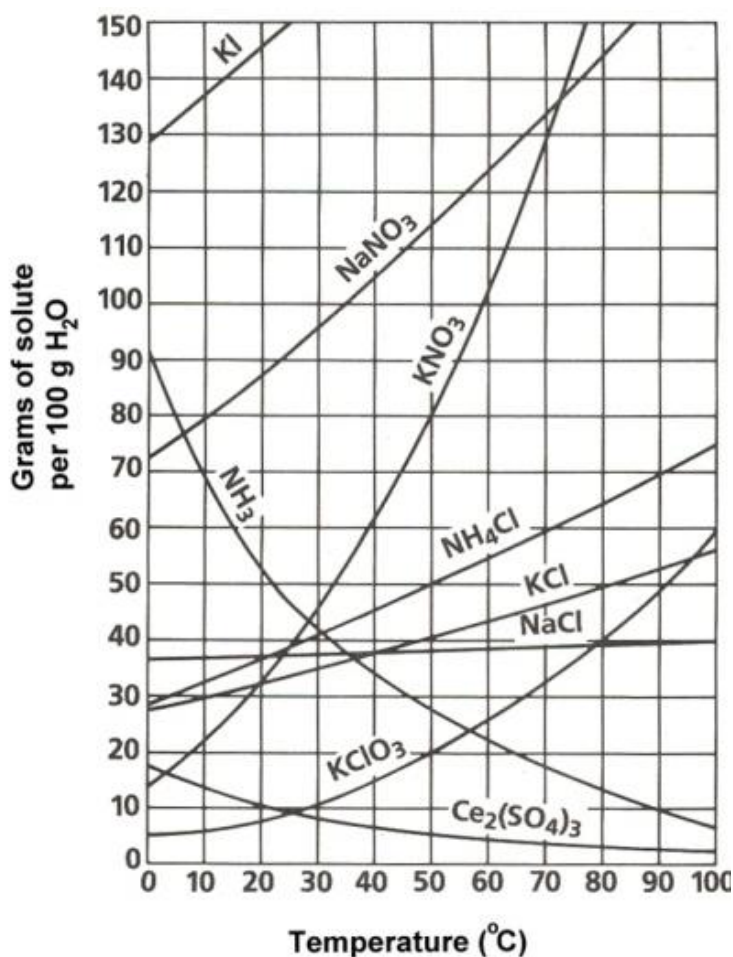
### **USING SOLUBILITY CURVES...**

- Any amount of solute below the line indicates the solution is unsaturated at a certain temperature.
- Any amount of solute above the line in which all of the solute has dissolved shows the solution is supersaturated.
- If the amount of solute is above the line but has not all dissolved, the solution is saturated and the extra grams of solute have precipitated (settled) to the bottom. (# g precipitated = total # g in solution - # g of a saturated solution at that temperature)
- Problems using an amount of solvent other than the amount given on the graph (usually 100 g or mL) can be solved by using proportions. See the following example...

**EX: How many grams of KClO<sub>3</sub> are required to saturate 550. mL of water at 40.0°C?**

- Determine the number of grams required to saturate 100 g of water at 40°C by reading your graph. Remember that right on the line represents saturation. **ANSWER: 15 g**
- Since the density of water is 1.00 g/mL, we can assume the 550 mL of water is equal to 550 g of water.
- Set up a proportionality to determine the amount of salt per 550 g:

$$\frac{15 \text{ g KClO}_3}{100 \text{ g H}_2\text{O}} = \frac{x}{550 \text{ g H}_2\text{O}} \qquad x = 82.5 \text{ g of KClO}_3$$



**EXAMPLE 12-15.** (example 12-10 on video) Are the following solutions saturated, unsaturated or supersaturated?

40. g of KCl completely dissolved in 100 g of water at 80.°C
120. g of KNO<sub>3</sub> completely dissolved in 100 g of water at 60.°C
80. g of NaNO<sub>3</sub> completely dissolved in 100 g of water at 10.°C

**EXAMPLE 12-16.** (example 12-11 on video) Which of the salts shown on the graph is the least soluble in water at 10.°C?

**EXAMPLE 12-17.** (example 12-12 on video) Which of the salts shown on the graph has the greatest increase in solubility as the temperature increases from 30. degrees to 60. degrees?

**EXAMPLE 12-18.** (example 12-13 on video) Which of the salts has its solubility affected the least by a change in temperature?

**EXAMPLE 12-19.** (example 12-14 on video) At 20.0°C, a saturated solution of sodium nitrate contains 100. grams of solute in 100. mL of water. How many grams of sodium nitrate must be added to saturate the solution at 50.0°C?

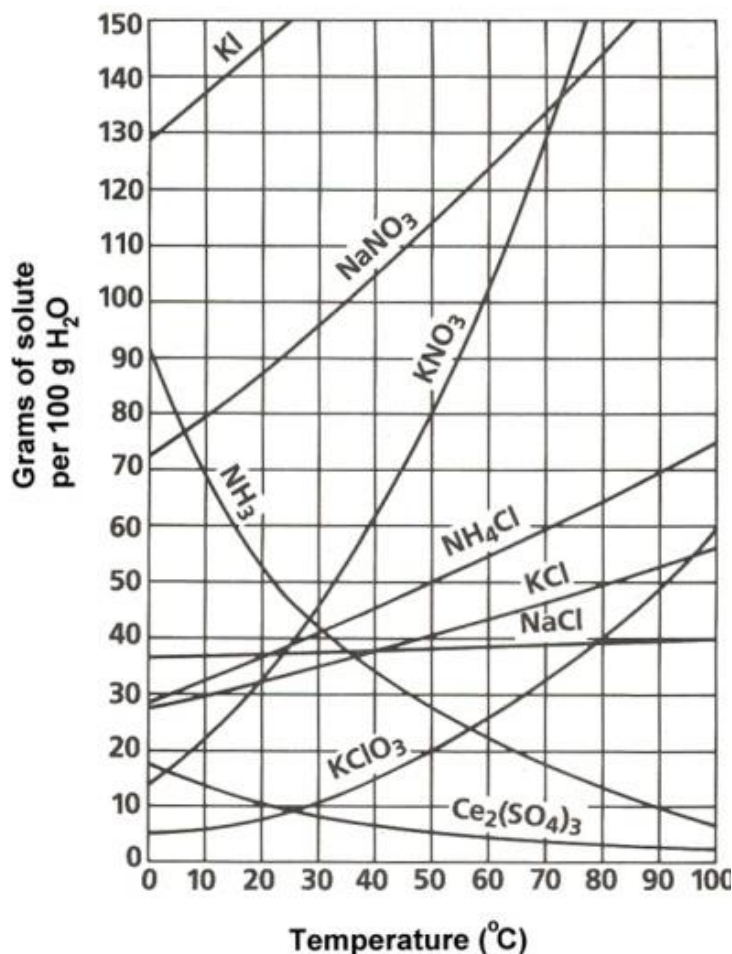
**EXAMPLE 12-20.** (example 12-15 on video) How many grams of potassium nitrate will it take to saturate 100. grams of water at 50.0°C?

**EXAMPLE 12-21.** (example 12-16 on video) How many grams of potassium nitrate would it take to saturate 300. grams of water at 50.0°C?

**EXAMPLE 12-22.** (example 12-17 on video) If 100. grams of water containing a saturated potassium nitrate solution at 50.0°C is cooled to 10.0°C, how many grams of potassium nitrate will precipitate out of the solution?



**EXAMPLE 12-23.** (example 12-18 on video) If a solution is made using 400. grams of water at 20.0°C and 40.0 grams of potassium chlorate, would the solution be saturated?



**EXAMPLE 12-24.** (example 12-19 on video) How many grams of potassium chlorate must be added to 1.00 liter of water to produce a saturated solution at 50.0°C?

**EXAMPLE 12-25.** (example 12-20 on video) If 50.0 mL of water saturated with KClO<sub>3</sub> at 25.0°C is slowly evaporated to dryness, how many grams of the dry salt would be recovered?

**EXAMPLE 12-26.** (example 12-21 on video) 30.0 grams of KCl are dissolved in 100. mL of water at 45.0°C. How many additional grams of KCl are needed to make the solution saturated at 80.0°C?

**EXAMPLE 12-27.** (example 12-22 on video) 80.0 grams of potassium nitrate are added to 100. grams of water at 10.0°C. How much of it will dissolve and how much will still be solid on the bottom of the beaker?

**EXAMPLE 12-28.** (example 12-23 on video) To what temperature Celsius must you raise the 100. grams of water in order for all 80.0 grams of the potassium nitrate to dissolve?

**EXAMPLE 12-29.** (example 12-24 on video) Suppose you had 160. grams of potassium nitrate and added it to 200. grams of water at 10.0°C. How much of it will dissolve and how much will still be solid on the bottom of beaker?



**EXAMPLE 12-30.** To what temperature Celsius must you raise the 200. grams of water in order for all 160. grams of the potassium nitrate to dissolve?

**EXAMPLE 12-31.** If 600. grams of water at 80.0°C are saturated with sodium chloride and then the solution is cooled to 20.0°C, how many grams of salt would you expect to “fall out” of solution?

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<http://youtu.be/B78Ffkavlo>

<https://vimeo.com/61856961>

## IX. USING CONCENTRATION

\_\_\_\_\_ : The amount of a substance per defined space. On the solubility curves, this was expressed in terms of grams of solute per 100 grams of solvent.

When we begin to use solubility in working problems, we rarely express the concentration of a solution in terms of grams of solute per 100 grams of solvent like the solubility curves do. Rather, we use the concept of Molarity.

A. **MOLARITY:** the number of \_\_\_\_\_ of solute per \_\_\_\_\_.

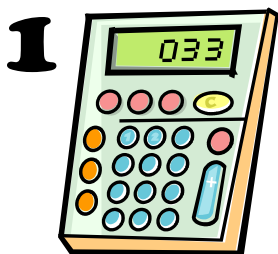
$$M = \frac{n_{\text{solute}}}{L_{\text{solution}}} \qquad M = \frac{\text{mass} / \text{MM}}{L_{\text{solution}}}$$

**UNITS: mol/L or M** (be familiar with both!)

Example: If 2.0 moles of potassium nitrate is dissolved in enough water to make 1.0 liter of total solution, we would say that the concentration of the solution is 2.0 M or 2.0 mol/L, or we could also say that we have a 2.0 Molar solution.

Molarity is the most common way of expressing concentration of solutions. An increase in molarity means that the solution has become \_\_\_\_\_.

**How would you prepare 1.0 liter of a 2.0 Molar solution of potassium nitrate? You need a VOLUMETRIC FLASK to prepare the solution because a volume is involved!**



1. Calculate how many grams 2.0 moles of potassium nitrate would be.
2. Mass out that many grams of potassium nitrate on a balance.
3. Add the amount of potassium nitrate you massed out to a 1.0 liter volumetric flask.
4. Add distilled water until the bottom of the meniscus is at the line and the total amount of solution is 1.0 liter. (Note: You will not be adding an entire 1.0 liter of water because the 2.0 moles of potassium nitrate will take up some space when dissolved.)

**EXAMPLE 12-32.** (12-27 on video) How many grams of sodium hydroxide must be used to make 500. ml of a 2.5 M sodium hydroxide solution?

How would you go about preparing this solution in the laboratory?

**EXAMPLE 12-33.** (12-28 on video) If 35.8 grams of lithium hydroxide are dissolved in enough water to make 750 ml of solution, what is the molar concentration (molarity) of the solution?

**EXAMPLE 12-34.** (12-29 on video) If 40.0 grams of sodium chloride are used to make a solution whose molarity is known to be 0.50 M, how many milliliters of solution do you have?

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<http://youtu.be/sLNuBL9rC5k>

<https://vimeo.com/38447755>

## B. REACTION STOICHIOMETRY USING MOLARITY

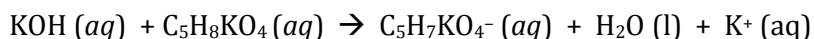
Because molarity involves moles, we can use molarity in the context of stoichiometric problems. Molarity can either be calculated (by finding moles using stoichiometry, and dividing it by liters), or it can be used as a conversion factor (where a molarity of “# mol/L” can be interpreted as a conversion of “# mol = 1 L”).

**EXAMPLE 12-35.** (12-45 on video) How many milliliters of 0.150 M Potassium Iodide solution would be required to react completely with 15.00 grams of Lead (II) Nitrate?

**EXAMPLE 12-36.** Using the same reaction as the previous example, how many grams of Lead (II) Iodide could potentially be formed if 65.0 mL of 0.550 M Potassium Iodide solution were reacted with excess Lead (II) Nitrate?

**EXAMPLE 12-37.** (note numbering mistake on video) If 25.00 mL of Hydrochloric Acid with a concentration of 0.1234 M is neutralized by 23.45 mL of Calcium Hydroxide, what is the concentration of the Calcium Hydroxide base?

**EXAMPLE 12-38.** What is the molarity of an aqueous solution of potassium hydroxide if 40.0 mL of the solution was fully neutralized by 0.230 g of potassium hydrogen phthalate (MM = 204.22 g/mol), as shown in the reaction below?



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<http://youtu.be/swe7lziEgU>

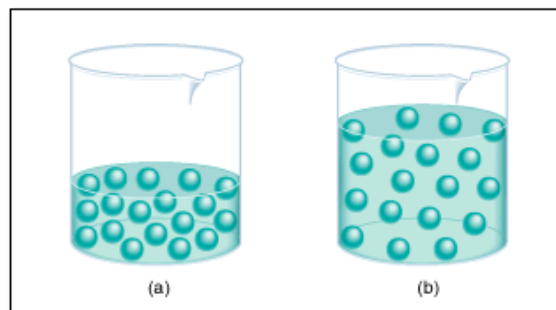
<https://vimeo.com/38434822>

## X. DILUTION

Dilution is the process of making a solution \_\_\_\_\_. This is usually done by adding more solvent to the solution.

**EXAMPLE 12-39.** (video 12-33) Which of these solutions is more concentrated?

**EXAMPLE 12-40.** (video 12-34) Which of these solutions is more dilute?



We usually use Molarity in dilution problems, as it's the most common way to express concentration. When we add more solvent, we \_\_\_\_\_ the concentration of the solution, which \_\_\_\_\_ molarity of the solution, and makes the solution more \_\_\_\_\_.

$$V_1M_1 = V_2M_2$$

vroom-vroom!



**EXAMPLE 12-41.** (video 12-35) What is the molarity of a solution of  $\text{NH}_4\text{Cl}$  prepared by diluting 50.00 mL of a 3.79 M  $\text{NH}_4\text{Cl}$  solution to 2.00 L?

**EXAMPLE 12-42.** (video 12-36) A 1.0 L, 3.8 M solution of  $\text{FeSO}_4$  has 7.0 L of water poured into it. What is the molarity of the new diluted solution?

## XI. COLLIGATIVE PROPERTIES

When you place a solute into a solvent, some of the solvent's properties can change. The properties that depend on the amount of solute present are called colligative properties.

The two colligative properties that you need to be familiar with are...

1. **Boiling Point** \_\_\_\_\_

2. **Freezing Point** \_\_\_\_\_

What this means is that a solution made of a solid or liquid solute dissolved in a liquid solvent will boil at a temperature \_\_\_\_\_ than the boiling temperature of the pure solvent, and will freeze at a temperature \_\_\_\_\_ than the freezing temperature of the pure solvent.

This explains why we can use salt on our sidewalks on ice days! It will cause the freezing point to lower, thereby causing the ice to melt.

This also explains why when cooking, some people will put salt in a pot of boiling water. It will cause the boiling point to elevate, meaning the temperature of the water will be higher than  $100^\circ\text{C}$ , and so your pasta will cook faster!

The more particles a substance forms when it dissolves, or the higher concentration of solute present, the more change one will in colligative properties.

**EXAMPLE 12-43.** Would a substance with  $i = 3$  or  $i = 4$  create more of a change in the freezing point and boiling point of a solvent? Explain.

**EXAMPLE 12-44.**  $\text{MgCl}_2$  is used more frequently to treat icy streets than  $\text{NaCl}$  is used. What could be a reason behind this?

**EXAMPLE 12-45.** Would 50.0 mL of 0.500 M glucose solution ( $\text{C}_6\text{H}_{12}\text{O}_6$ ) or 50.0 mL of 0.500 M ammonium hydroxide solution ( $\text{NH}_4\text{OH}$ ) cause more of a change in the freezing point and boiling point of 100 mL of water? Explain.

