

**Ms. Randall**  
**Regents Chemistry**  
**Unit 8 Solutions**

**Unit Objectives:**

- Differentiate between heterogeneous and homogeneous mixtures
- Identify the various types of heterogeneous mixtures and their properties
  - colloid
  - suspension
- Identify the various types of homogeneous mixtures and their properties
  - alloy
  - solution
- Define solubility and understand the factors that contribute to solubility nature of solute and solvent (like dissolves like), temperature
- Use Table F to determine if precipitate is formed in a chemical reaction
- Distinguish between saturated, unsaturated, or supersaturated solutions
- Read the solubility curve (Table G) to determine if a solution is (1) saturated, (2) unsaturated, or (3) supersaturated
- Differentiate between dilute and concentrated solutions
- Calculate various concentrations of a solution using the following:
  - Molarity (M)
  - Percent by Mass
  - Percent by Volume
  - Parts per Million (ppm)
- Prepare a solution of known concentration
- Explain a solute's effect on a solution (colligative properties)
  - Freezing Point Depression
  - Boiling Point Elevation
  -

***Focus Questions for the Unit:***

- **How can precise concentrations be made?**
- **How are electrolytes described, in terms of formula, properties, concentration?**

**YOU SHOULD BE ABLE TO ANSWER THESE IN DETAIL BY THE END OF THE UNIT**

**Define the following vocabulary:**

Substance

Homogeneous mixture

Heterogeneous mixture

Solution

Solute

Solvent

Insoluble

Immiscible

Soluble

Solubility

Chromatography

Distillation

Concentration

Molarity

Colligative Properties

Saturated Solution

Unsaturated Solution

Supersaturated Solution

Precipitate

Dilution

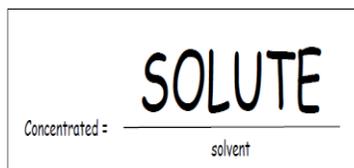
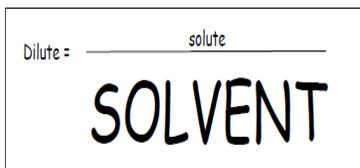
Alloy

## Lesson 1: Types of Solutions and Solubility curves

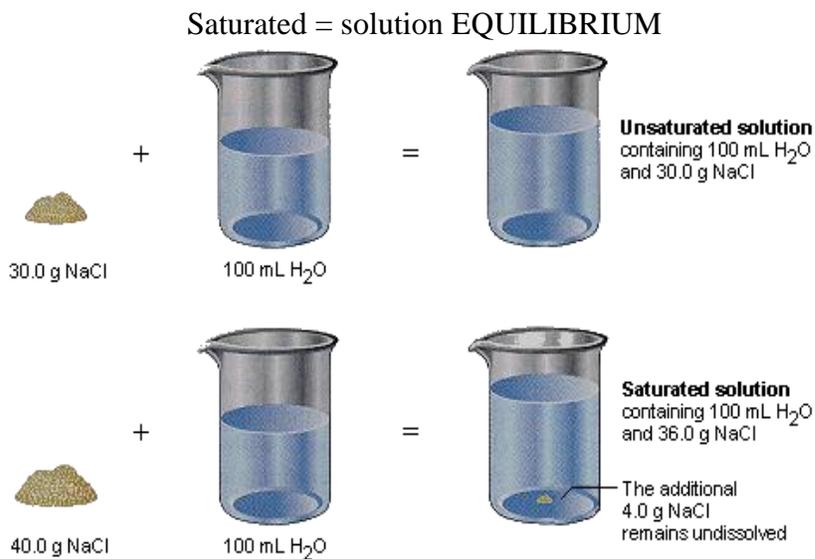
### *Objective:*

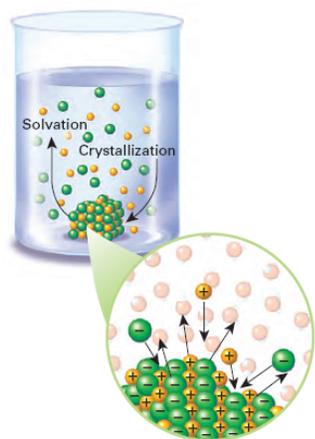
- *Determine if a solution is saturated, unsaturated or supersaturated using Table G*
- *Determine how much of a solute can dissolve, will precipitate etc. using Table G*

**A. Solubility** is a term used to describe the amount of solute which can be dissolved in a solvent. The solubility of two substances will depend on the similarity of the two substances. A **concentrated solution** is one where the amount of solute in a given volume of solvent is great. A **dilute solution** is one where the amount of solute in a given volume of solvent is relatively small.



1. **UNSATURATED:** a solution in which MORE SOLUTE could be dissolved in a given volume of solvent
2. **SATURATED:** a solution containing the MAXIMUM AMOUNT of solute that will dissolve in a given volume of solvent



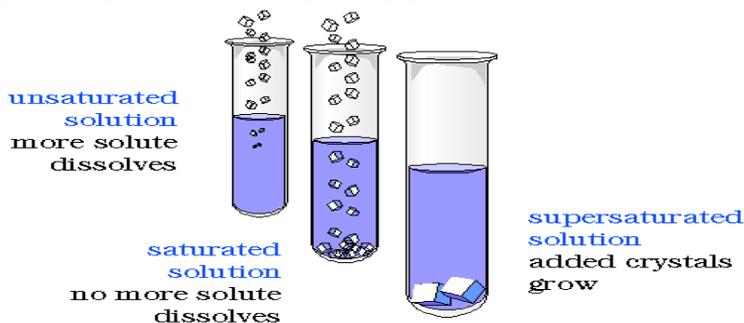


**Figure 16.2** In a saturated solution, a state of dynamic equilibrium exists between the solution and the excess solute. The rate of solvation (dissolving) equals the rate of crystallization, so the total amount of dissolved solute remains constant.  
**Inferring** What would happen if you added more solute?

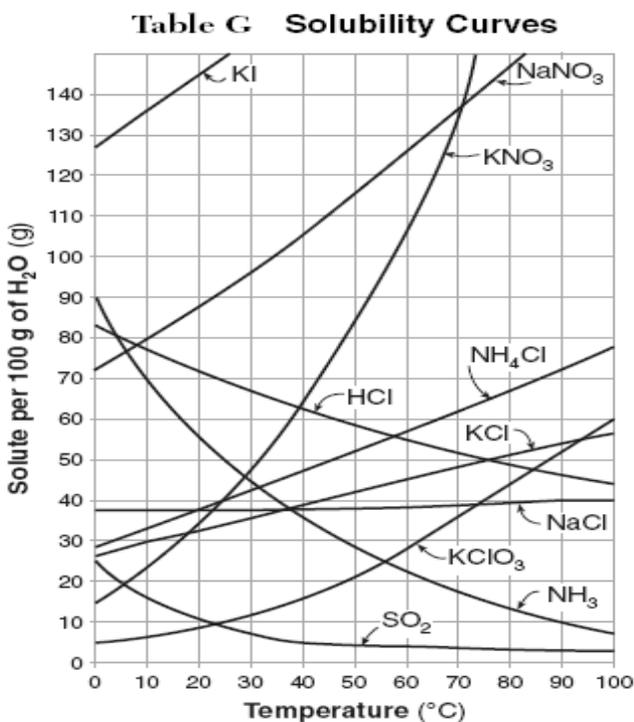
3. **SUPERSATURATED:** a solution that contains MORE SOLUTE than would normally dissolve in a given volume of solvent; usually requires an INCREASE in TEMPERATURE or PRESSURE *initially* (VERY UNSTABLE)

\*\*\*\*\*How to tell if a solution is saturated/unsaturated or supersaturated...\*\*\*\*\*

### Saturated Solutions



### B. Solubility Curves



**HOW TO USE TABLE G:** Table G shows the number of grams of a substance that can be dissolved in **100 GRAMS OF WATER** at temperatures between 1°C and 100°C.

Each curve represents the **maximum amount** of a substance that can be **dissolved** at a **given temperature**

- A solution which is **on the curve** is SATURATED
- A solution **below the curve** is UNSATURATED
- A solution **above the curve** is SUPERSATURATED

**Example 1:** Adding solute to unsaturated solution: A solution of KCl contains 20.g of KCl in 100g of H<sub>2</sub>O at 30. C. How many more grams of KCl must be added to make it saturated?

Saturated amount = 35 g/100g H<sub>2</sub>O

Grams given in problem= 20.g

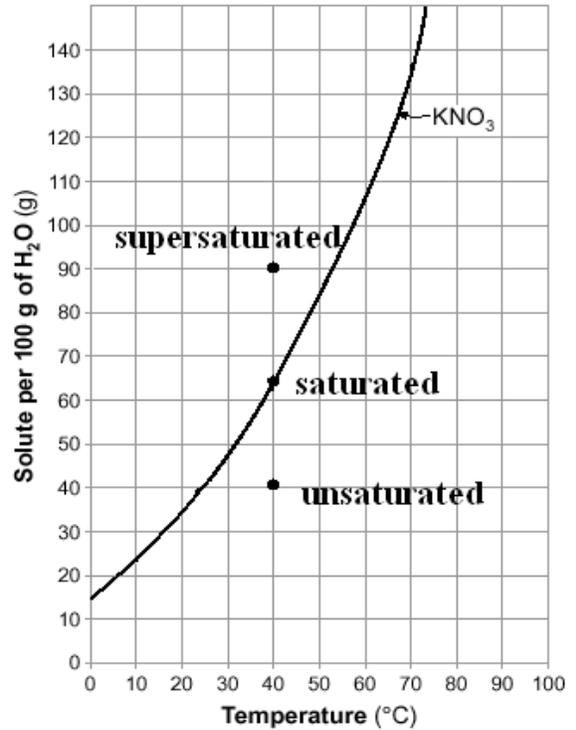
Amount needed to create a saturated solution = add 15 g of KCl

**Example 2:** Cooling a solution: A saturated solution of KNO<sub>3</sub> is cooled from 70. C to 25. C. How much of the KNO<sub>3</sub> will precipitate (become solid, settle out)?

Saturated amount at 70° C = 135g/100g H<sub>2</sub>O

Saturated amount at 25° C= 40g/100g H<sub>2</sub>O      Amount of  
KNO<sub>3</sub> will settle out. =95g of

**Table G Solubility Curves**

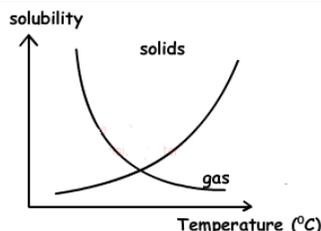


**Example 3:** Which salt is least soluble at 50. °C, NaCl or NH<sub>4</sub>Cl?

NaCl because less can be dissolved at that temperature than NH<sub>4</sub>Cl.

**C. Solubility of Solids Vs. Gases**

Solids	Gases
As temperature <b>INCREASES</b> , solubility of a solid <b>INCREASES</b>	As temperature <b>INCREASES</b> , the solubility of a gas <b>DECREASES</b>
	As pressure <b>INCREASES</b> , the solubility of a gas <b>INCREASES</b>



## Lesson 2: Factors affecting solubility

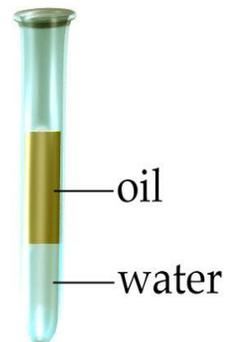
### **Objective:**

- *Identify the factors that affect solubility*

### **Watch this:** Factors Affecting Solubility

#### 1. Nature of SOLUTE & SOLVENT(Polarity)

- “**Like**” substance DISSOLVE “**like**” substances
- EXAMPLE: Ammonia (polar) dissolves in water (polar)
  
- Oil (nonpolar) DOES NOT dissolve in water (polar)
- Salt (ionic) DISSLOVES in water (polar)
- 



#### 2. Temperature

- For solids, as temperature  $\uparrow$ , solubility  $\uparrow$

**Remember:** Table G is there to help you!

Gases slope  $\downarrow$  solids slope  $\uparrow$ .

- For MOST gases, as temperature  $\uparrow$ , solubility  $\downarrow$   
Example: heat soda  $\rightarrow$  carbonation is LOST!

#### 3. Pressure

Solid/Liquid Solutes	Gases Dissolved in Liquids
LITTLE TO NO EFFECT	As pressure increases, VOLUME $\downarrow$ and SOLUBILITY $\uparrow$

Recall: We can change the volume of gases!

Reason: Gas is forced into less space per liquid, so it DISSOLVES more!

Ex: can of soda

### Factors affecting rate of dissolving

The **rate of solution** is a measure of how fast a substance dissolves.

1. **Size of the particles (surface area)** -- When a solute dissolves, the action takes place only at the surface of each particle. When the total surface area of the solute particles is increased, the solute dissolves more rapidly. *Breaking a solute into smaller pieces increases its surface area and increases its rate of solution.*
2. **Stirring** -- With liquid and solid solutes, stirring brings fresh portions of the solvent in contact with the solute. Stirring, therefore, allows the solute to dissolve faster.

### Lesson 3: Solubility of Ionic Compounds

#### *Objective:*

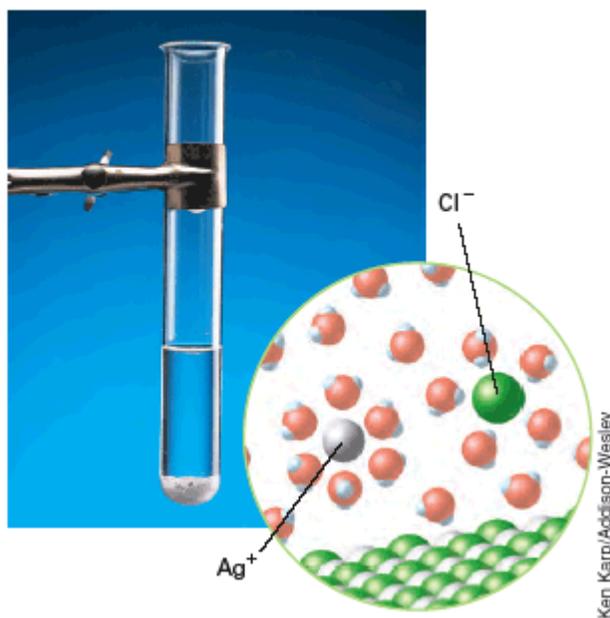
- *Use Table F to determine solubility*
- *Compose double replacement reactions and determine the precipitate*

Not all ionic compounds can dissolve or are **soluble** in water. In order to decide, you will learn how to use Table F!! Take a look now! Substances that form in water stay dissolved because they ARE soluble in water, while others form “**precipitates**” and settle out as a solid because they ARE NOT soluble.

Chemists use the word “**dissociate**” to describe the falling apart process that a soluble ionic compound undergoes. In other words, the ions in the crystal are pulled apart by the action of water molecules. Water molecules are able to pull ions into solution by pulling on them with electrical attractions. The oxygen end of a water molecule has a somewhat negative charge to it and thus will be attracted to the positive ions in the salt crystal. Conversely, the hydrogen side of water molecules has a somewhat positive charge to it, and is attracted to negative ions in the crystal. Water is a polar molecule.

The process of water molecules surrounding the salt and dissolving it one ion at a time is called “**hydration.**” The process of removing water to make something return to a solid is called “**dehydration.**”

You should learn how to write a “dissociation equation.” Here are a couple examples:



**Figure 18.16** Silver chloride is slightly soluble in water.

“Dissociate” is a shortened up form of “dis-associate”. If an ionic compound dissolves in water, it is because the attraction of water molecules for the ions is able to pull the ions apart from each other, overcoming their ionic bonds and causing them to “dis-associate” from each other. For ionic compounds that dissolve in water, describing them as “**electrolytes**” is appropriate, since the crystal will fall apart in water. **An electrolyte is a substance that dissolves in water and forms a solution capable of conducting an electric current.** The ability of a solution to conduct an electric current depends upon the concentration of ions that are present.

**Ionic compounds that do not dissolve** are said to be **insoluble**.

Reference **TABLE F** lists solubility rules and exceptions to those rules.

- ✓ **Precipitate = INSOLUBLE product(solid)** that falls to the bottom of a mixture/solution following a double replacement reaction
- ✓ Precipitate will form (or a visible reaction will occur) if ONE or BOTH of the products listed is INSOLUBLE

### Using Table F

**Find one of the negative ions(anions)** in your formula in either **Column 1 or Column 3** (there are only two positive ions listed in Table F, so it is more likely that you will find the “ion in the back” of your formula before the “ion in the front”

- ❖ The title of the column tells you if the compound is SOLUBLE or INSOLUBLE.
- ❖ Check to see if there are any exceptions. If your compound contains one of the exceptions then it negates (opposite) the title of the column.

Ions That Form Soluble Compounds	Exceptions
Group 1 ions (Li <sup>+</sup> , Na <sup>+</sup> , etc.)	
ammonium (NH <sub>4</sub> <sup>+</sup> )	
nitrate (NO <sub>3</sub> <sup>-</sup> )	
acetate (C <sub>2</sub> H <sub>3</sub> O <sub>2</sub> <sup>-</sup> or CH <sub>3</sub> COO <sup>-</sup> )	
hydrogen carbonate (HCO <sub>3</sub> <sup>-</sup> )	
chlorate (ClO <sub>3</sub> <sup>-</sup> )	
perchlorate (ClO <sub>4</sub> <sup>-</sup> )	
halides (Cl <sup>-</sup> , Br <sup>-</sup> , I <sup>-</sup> )	when combined with Ag <sup>+</sup> , Pb <sup>2+</sup> , and Hg <sub>2</sub> <sup>2+</sup>
sulfates (SO <sub>4</sub> <sup>2-</sup> )	when combined with Ag <sup>+</sup> , Ca <sup>2+</sup> , Sr <sup>2+</sup> , Ba <sup>2+</sup> , and Pb <sup>2+</sup>
Ions That Form Insoluble Compounds	Exceptions
carbonate (CO <sub>3</sub> <sup>2-</sup> )	when combined with Group 1 ions or ammonium (NH <sub>4</sub> <sup>+</sup> )
chromate (CrO <sub>4</sub> <sup>2-</sup> )	when combined with Group 1 ions or ammonium (NH <sub>4</sub> <sup>+</sup> )
phosphate (PO <sub>4</sub> <sup>3-</sup> )	when combined with Group 1 ions or ammonium (NH <sub>4</sub> <sup>+</sup> )
sulfide (S <sup>2-</sup> )	when combined with Group 1 ions or ammonium (NH <sub>4</sub> <sup>+</sup> )
hydroxide (OH <sup>-</sup> )	when combined with Group 1 ions, Ca <sup>2+</sup> , Ba <sup>2+</sup> , or Sr <sup>2+</sup>

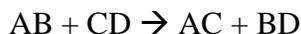
\*\*\*\*\*The “exceptions” are the opposite. \*\*\*\*\*

**EXAMPLE:** Is the compound AgCl soluble or insoluble?

The Cl<sup>-</sup> is found in the soluble column BUT when it is bonded to Ag<sup>+</sup> it is an exception so the compound is INSOLUBLE.

**Application:** Double replacement reactions and Table F.

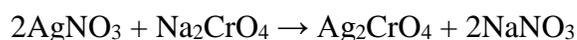
Recall: When two aqueous IONIC compounds are mixed together ( switch partners) you have a **double replacement reaction**.



One of the products forms a solid (precipitate) -Table F tells you whether products are soluble or insoluble

**EXAMPLE:** Silver nitrate and sodium chromate solutions are mixed together. Will a precipitate form? If so, what is the name of the precipitate?

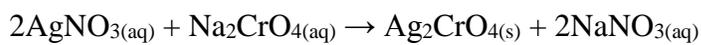
STEP 1 → Write the chemical or word equation for the **double replacement** reaction. Remember that you must look at oxidation states when determining the formula for an ionic compound! Don't forget to look to table E for those polyatomic ions too! Make sure to balance your equation!



STEP 2 → Check the solubility for both products (on Table F). If there are any that are insoluble according to Table F, you have found your precipitate (s)!

$Ag_2CrO_{4(s)}$  = precipitate.

STEP 3 → Now write in your states of matter(punctuation)!



### **Using Table F to determine concentration:**

CONCENTRATION is directly related to amount of solute DISSOLVED. When looking to find concentration in terms of solubility, use reference Table F.

- Most concentrated = SOLUBLE compounds
- Least concentrated = INSOLUBLE compounds

***\*If a compound doesn't dissolve, it doesn't add to the concentration of the solution!***

## Lesson 4: Concentration

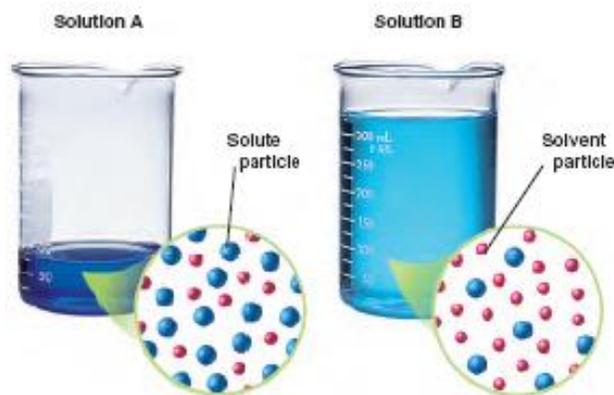
### *Objective:*

- *Calculate the concentration of various solutions*

**Concentration** is a measurement of the amount of a solute that is dissolved in a given amount of a solvent. Two **qualitative** ways of describing the concentration of solutions are to use the terms “dilute” and “concentrated”.

**Dilute** solutions are solutions which have a small amount of solute dissolved as compared to the amount of solvent. **Concentrated** solutions have a lot of solute dissolved as compared to the amount of solvent.

Look at the Picture below:



Which of the two solutions is **more concentrated** with solute particles? How do you know? Think about it!

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The terms dilute and concentrated are scientifically **inexact** terms. Scientists like to be very exact or precise with their measurements. In other words, we like to be very close to the actual numerical value. These terms do not give us numbers so scientists use a calculation to find the concentration of particles. The formulas for concentration calculations can all be found on reference Table T.

1. Molarity (M)
2. Parts per Million (ppm)
3. Percent by Mass
4. Percent by Volume

### 1. **Molarity(M)**

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

Example: A solution has a volume of 2.5 liters and contains .70 mol of NaCl. What is the molarity?

$$M = \frac{.70 \text{ mol}}{2.5 \text{ L}} = 0.28M$$

**Note:** If a problem gives you grams instead of moles you must use the mole formula on table T first!  
Volume is always in Liters so if you are given mL you must convert the volume to liters.

## 2. Parts per million(PPM)

$$\text{parts per million} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 1000000$$

Example: A sample of water is found to contain 0.010 g lead in 15. g solution. What is the concentration in ppm?

$$\text{ppm} = \frac{0.010\text{g}}{15.\text{g}} \times 1,000,000 = 670 \text{ ppm}$$

**Note:** You may be given problems in which you have to solve for the mass of solute or solution. Use your formula carefully!

## 3. Percent by Mass(%)

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$$\% \text{ composition by mass} = \frac{\text{mass of part}}{\text{mass of whole}} \times 100$$

Example: What is the percent by mass of glucose if 2.8 grams of glucose is dissolved in 100g of water?

$$\% \text{ mass} = \frac{2.8 \text{ g glucose}}{100 \text{ g water}} = 0.28 \% \text{ glucose}$$

## 4. Percent by Volume

$$\%v/v = \frac{\text{Volume Solute}}{\text{Volume Solution}} \times 100$$

Example: What is the percent by volume of alcohol if 40.0 mL of acetone is diluted with water to form a total volume of 300. mL?

$$\frac{40.0 \text{ mL}}{300. \text{ mL}} \times 100 = 13.3 \%$$

## Lesson 5: Colligative Properties of Solutions

### *Objective:*

- *Differentiate between boiling point elevation and freezing point depression and the factors that influence them*

As we have discussed, solutions have different properties than either the solutes or the solvent used to make the solution. Those properties can be divided into two main groups--colligative and non-colligative properties.

**Colligative properties** depend only on the number of dissolved particles in solution and not on their identity.

**Non-colligative properties** depend on the identity of the dissolved species and the solvent.

To explain the difference between the two sets of solution properties, we will compare the properties of a 1.0 *M* aqueous sugar solution to a 0.5 *M* solution of table salt (NaCl) in water. Despite the concentration of sodium chloride being half of the sucrose concentration, both solutions have precisely the same number of dissolved particles because each sodium chloride unit creates two particles upon dissociation--a sodium ion, Na<sup>+</sup>, and a chloride ion, Cl<sup>-</sup>. Therefore, any difference in the properties of those two solutions is due to a non-colligative property. Both solutions have the same freezing point, boiling point and vapor pressure because those colligative properties of a solution only depend on the number of dissolved particles. The taste of the two solutions, however, is markedly different. The sugar solution is sweet and the salt solution tastes salty. Therefore, the taste of the solution is not a colligative property. Another non-colligative property is the color of a solution. A 0.5 *M* solution of CuSO<sub>4</sub> is bright blue in contrast to the colorless salt and sugar solutions. Other non-colligative properties include viscosity, surface tension, and solubility.

**Colligative Properties**-*Physical properties of solutions that depend on the concentration of solute in a given amount of solvent.*

When a solute is dissolved in a solvent, solvent molecules SURROUND the particles of the solute. The result...the BOILING POINT and FREEZING POINT of the solution to CHANGE in a very specific and predictable way. **Dissolved solute interferes with the intermolecular forces of attraction** which is why boiling and freezing points are affected.

**Boiling Point Elevation** = Boiling point. INCREASES when solute is added. The

- Ex: Adding salt to water allows you to boil pasta at 102-103°C (cooks it faster)

**Freezing Point Depression** = Freezing point DECREASES when solute is added

- Ex: Putting salt on roads causes ice to melt because it drops f.p. below 0°C.

## What factors affect colligative properties?

### 1. Electrolytes = Ionic compounds in solution are called electrolytes

The MORE IONS a solute breaks up into, the higher the boiling point and the lower the freezing point of the solution will be.



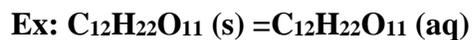
*One mole of sodium chloride dissolves to form one mole of sodium ions and one mole of chloride ions (2 moles of dissolved ions are produced).*



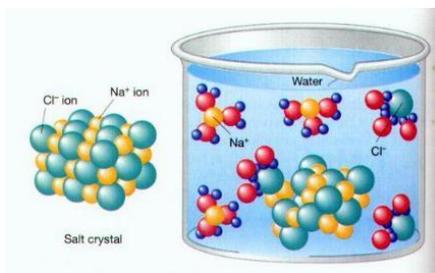
*One mole of calcium bromide dissolves to form one mole of calcium ions and two moles of bromide ions (3 moles of dissolved ions are produced)*

### 2. Nonelectrolytes (covalent compounds)

Less impact on the melting and boiling point of a solution than ionic compounds do  
They DO NOT break up any further in solution



*One mole of sucrose dissolves to form one mole of dissolved sucrose.*



### IONIC VS MOLECULAR (COVALENT) SOLUTES:

	MOLECULAR: C <sub>12</sub> H <sub>22</sub> O <sub>11</sub> (Sucrose – covalent!)	IONIC: NaCl
Reaction produced when solute dissolved in water	$\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s}) \xrightarrow{\text{H}_2\text{O}} \text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{aq})$	$\text{NaCl}(\text{s}) \xrightarrow{\text{H}_2\text{O}} \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
Number of moles of product produced	<b>1 mol reactants ⇌ 1 mol products</b>	<b>1 mol reactants ⇌ 2 mol products</b>
General Rule	The solute that dissolves to form a greater number of products (ions) will have a greater f.p. depression (f.p. will ↓ by more) and a greater boiling point elevation (b.p. will ↑ by more)	
Example:	$\text{CaCl}_2(\text{s}) \xrightarrow{\text{H}_2\text{O}} \text{Ca}^+(\text{aq}) + 2\text{Cl}^-(\text{aq})$	
# ions produced:	<b>3 moles of ions produced</b>	

\*

Why do we salt the roads in the winter rather than sugar them?!?!? Freezing point depressions:



*\*It's always a good idea to have a little CaCl<sub>2</sub> on hand for when the REALLY cold weather hits...*