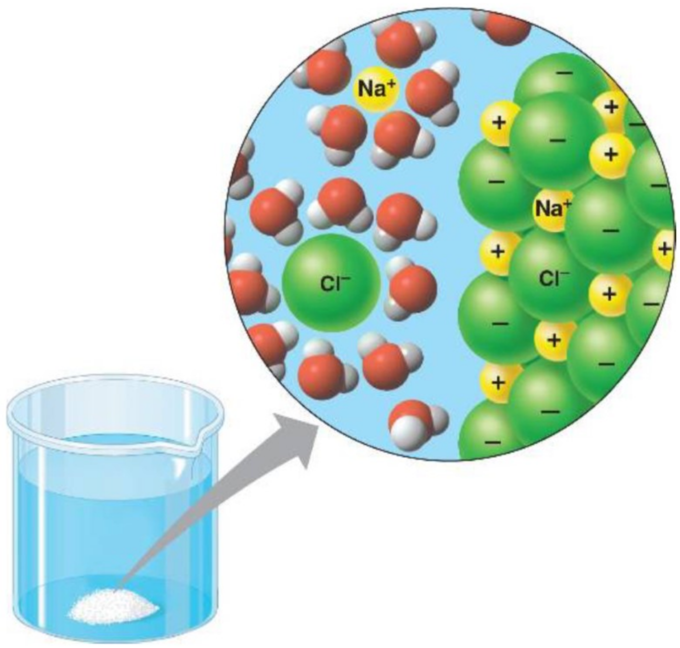
Note Packet
UNIT 7: SOLUTIONS



Unit Vocabulary:

Alloy	heterogeneous	saturated
Aqueous	homogeneous	solubility
boiling point	insoluble	soluble
boiling point elevation	miscible	solution
colligative properties	mixture	solute
colloid	molarity (M)	solvent
concentrated	parts per million (ppm)	supersaturated
concentration	percent by mass	suspension
dilute	percent by volume	Tyndall Effect
freezing point depression	precipitate	unsaturated

Unit Objectives (Upon the Completion of this unit you will be able to):

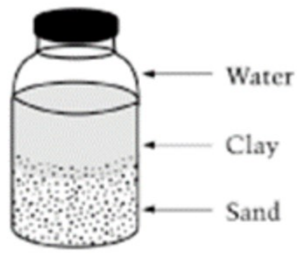
- Differentiate between heterogeneous and homogeneous mixtures
- Identify the various types of heterogeneous mixtures and their properties
  - colloid
  - suspension
- Identify the various types of homogeneous mixtures and their properties
  - alloy
  - solution
- Define solubility and understand the 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

A mixture is: a combination of 2+ pure substances (elements and/or compounds) that can be separated by PHYSICAL means

Mixtures can be categorized into one of two groups:

1. **HETEROGENEOUS MIXTURE:** mixture of 2 or more pure substances in DIFFERENT PHYSICAL STATES; VARYING COMPOSITION (***not*** uniform throughout)

- **Suspension:** heterogeneous mixture that ***clearly*** separates



- eventually SEPARATES upon standing
- solid does not visibly dissolve in the liquid (or break down)
- separate by FILTRATION (using a filter)
- relatively LARGE particle size
- Examples: SAND IN WATER, ITALIAN DRESSING

- **Colloid:** heterogeneous mixture that does NOT CLEARLY SEPARATE

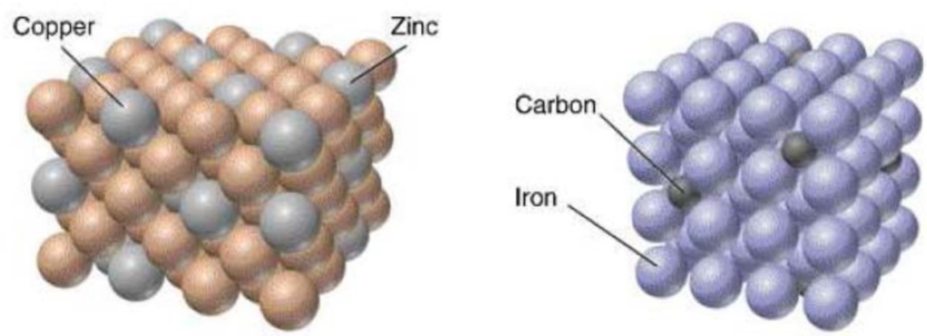
- CLOUDY (*appear* homeogenous, but are NOT)
- do not settle (naturally; on their own) through time
- can be separated by a CENTRIFUGE (forced gravity)

- MEDIUM to SMALL particle size
- properties in between a suspension and a solution
- Examples: BLOOD, MILK

2. HOMOGENEOUS MIXTURE (AKA Solution): mixture of 2 or more pure substances in the SAME PHYSICAL STATE; CONSTANT COMPOSITION (uniform throughout)

a. looks like CLEAR, COLORED or COLORLESS water
b. EXTREMELY SMALL particle size (you can't see them)
c. may exist in any phase (liquid phase most common)
d. Examples: SALTWATER, KOOL-AID, CLUB SODA

- **Alloy:** solution where at least one of the materials is a METAL (although most times it is a mixture of 2+ metals)



A Brass, a substitutional alloy **B** Carbon steel, an interstitial alloy

- Metals arrange in an alternating pattern with SEA OF MOBILE VALENCE ELECTRONS gluing them together

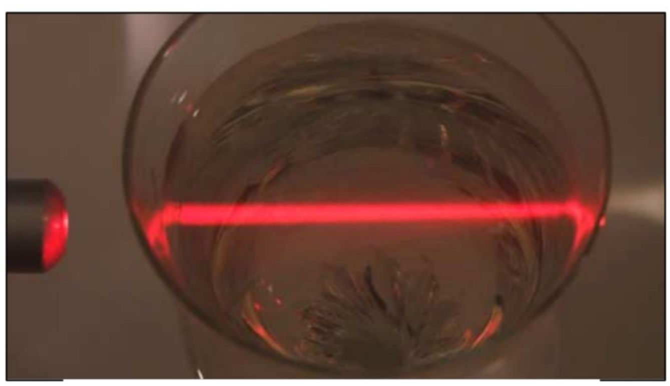
- Examples: BRASS (zinc + copper), STEEL (iron + carbon)

Now, we will focus on SOLUTIONS only…

Characteristics of SOLUTIONS

1. HOMOGENOUS mixtures
2. CLEAR (as opposed to cloudy) and do not disperse light (no Tyndall observed)
3. Can have COLOR or be COLORLESS
4. DO NOT SETTLE upon standing (after any amount of time)
5. CANNOT be separated by filtration

TYNDALL EFFECT: NOT seen in solutions; seen when a mixture makes a beam of light VISIBLE



(NOT a solution!)

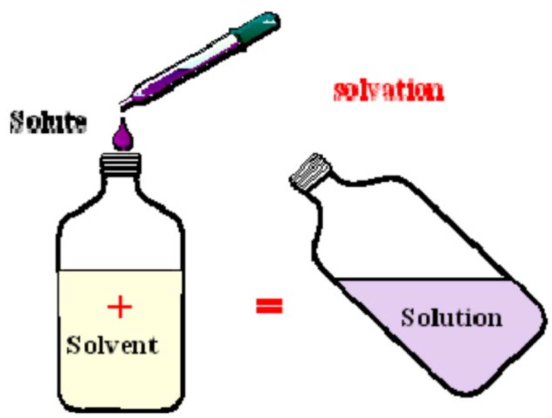
*****Solutions always assume the phase of the solvent!***

Phase of Solute	Phase of Solvent	Example
Solid	Solid	Metallic alloy
Solid	Liquid	Table salt & water
Liquid	Liquid	Ethyl alcohol & water
Gas	Liquid	CO_2 & water (club soda)
Gas	Gas	Air

SOLUTIONS ARE MADE UP OF 2 PARTS:

- ✓ **SOLUTE**: substance or substances that are DISSOLVED in a solution (example: the salt in saltwater); solute is usually a SOLID and is the LESSER quantity of the two (versus the solvent)

- ✓ **SOLVENT**: the substance that is DISSOLVING the solute (example: the water in salt water); solvent is usually a LIQUID and is the GREATER quantity of the two (versus the solute)

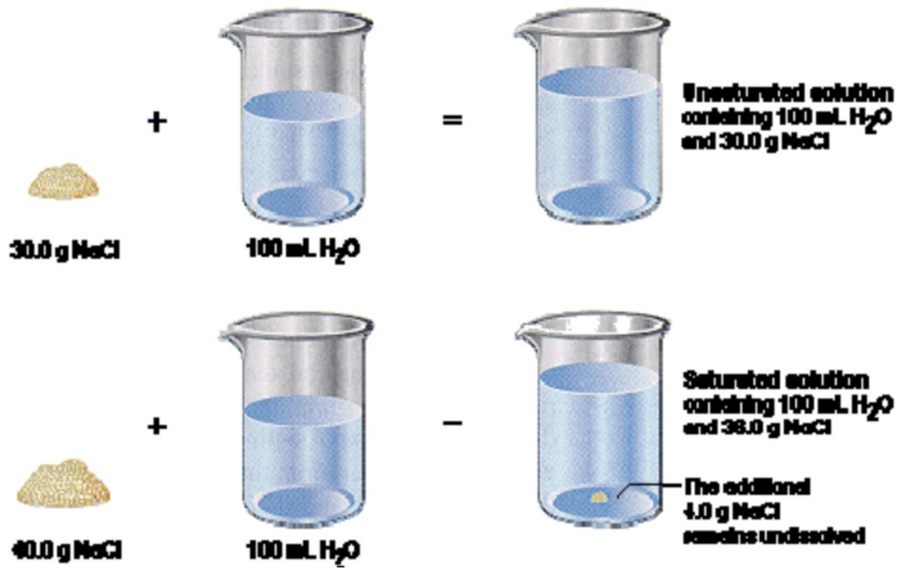


AQUEOUS = a solution in which WATER is the solvent (aq)

Example: NaCl(aq) is otherwise known as SALT WATER, is actually $Na^{+}(aq) + Cl^{-}(aq)$

DIFFERENT TYPES OF SOLUTIONS:

- ✓ **UNSATURATED**: a soln in which MORE SOLUTE could be dissolved in a given volume of solvent
- ✓ **SATURATED**: a soln containing the MAXIMUM AMT. of solute that will dissolve in a given volume of solvent
  - Saturated = solution EQUILIBRIUM

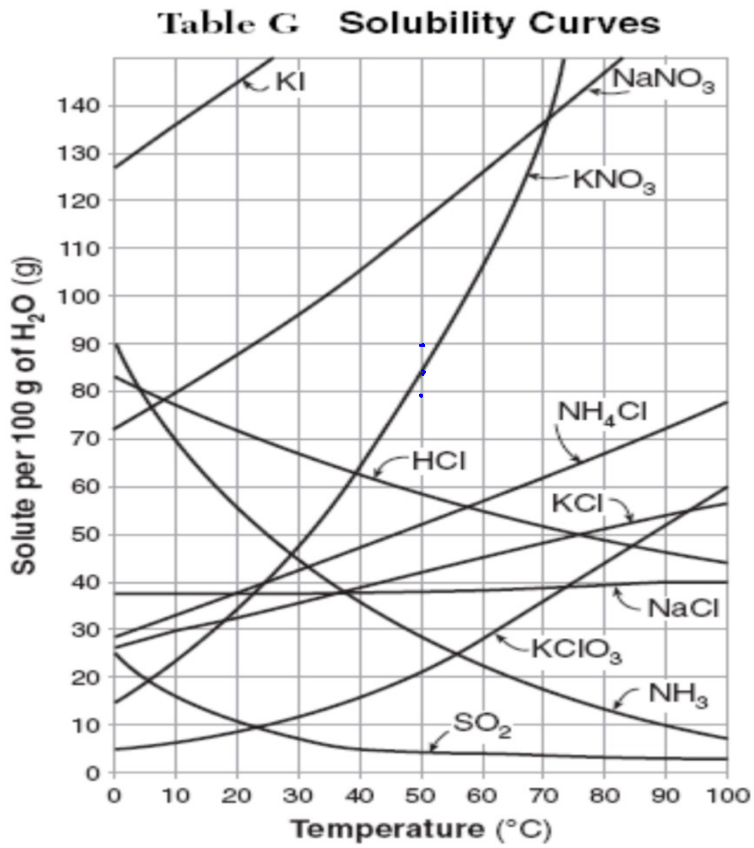


- ✓ **SUPERSATURATED**: a soln that contains MORE SOLUTE than would normally dissolve in a given volume of solvent; usually requires an INCREASE in TEMPERATURE or PRESSURE ***initially*** (VERY UNSTABLE)
  - ➢ *Very rare! Allow me to demonstrate…*

Situation:

- ➢ If you have an UNSATURATED solution and you add crystals, what will you observe? CRYSTALS DISSOLVE/DISAPPEAR
  - The rate of dissolving is GREATER than the rate of precipitation
- ➢ If you have a SUPERSATURATED solution and you add crystals, what will you observe? ALL CRYSTALS PRECIPITATE OUT
  - The rate of dissolving is LESS than the rate of precipitation
- ➢ If you have a SATURATED solution and you add crystals, what will you observe? *CRYSTALS SINK TO BOTTOM
  - *The rate of dissolving is EQUAL to the rate of precipitation

SOLUBILITY CURVES - USING TABLE G (of the Reference Tables):



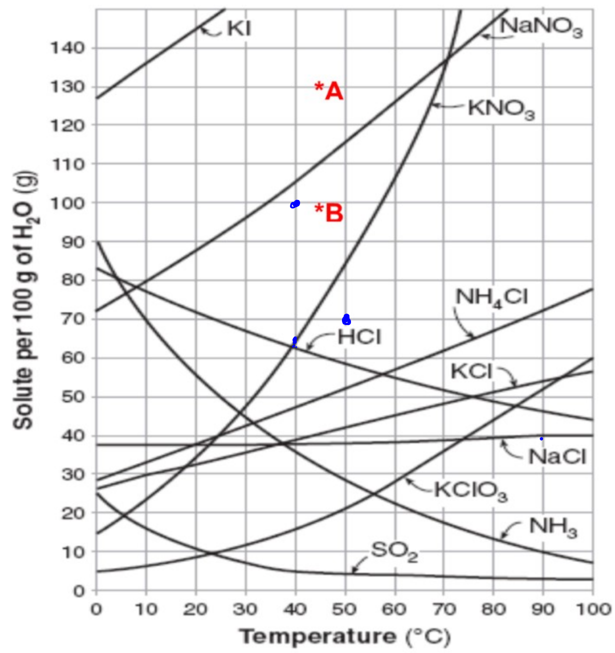
HOW TO USE TABLE G: Table G shows the number of grams of a substance that can be dissolved in **100 GRAMS OF WATER** at temperatures between 1°C and 100°C. Each curve represents the maximum amount of a substance that can be dissolved at a given temperature. All lines that show an increase in solubility as temperature increases represent **SOLIDS**. The lines for **GASES** slope downward.

*****If the intersection of GRAMS OF SOLUTE vs. TEMPERATURE falls:**

- **ON THE CURVE, solution is SATURATED**
- **UNDER THE CURVE, solution is UNSATURATED**
- **ABOVE THE CURVE, solution is SUPERSATURATED**

Use Table G to answer the following questions:

Table G Solubility Curves



1. How many grams of NaCl will dissolve in 100 g of H_2O at 90°C? 40g
2. How many grams of KCl will dissolve in **50 g** of H_2O at 30°C? 100g 38g ≈ 18g ± 1g
3. How many grams of KCl will dissolve in **200 g** of H_2O at 30°C? 276g ± 1g
4. Name two substances on Table G that are solids.
   KI & NaCl (slope up)
5. Name two substances on Table G that are gases.
   HCl & NH3 (slopes down)
6. Is a solution that is made up of 100 grams of H_2O and 70 grams of KNO_3 at 50°C saturated, unsaturated, or supersaturated? Why?
7. At 40°C, if 100 grams of KNO_3 are added to 100 grams of water, how many grams of solute will fall to the bottom of the solution (and won't dissolve)?
   100g − 65g = 35g will fall out of solution
8. Which is more concentrated on Table G, point A or point B? Why?
   Point A is more concentrated

SOLUBILITY FACTORS (Why do some things dissolve and others do not?)

Factor 1 – *Nature of Solute & Solvent*

	Nonpolar Solvent (Acetone)	Polar Solvent (Water)
Nail polish — Nonpolar Solute	SOLUBLE	INSOLUBLE
Sugar — Polar Solute	INSOLUBLE	SOLUBLE
Salt — *Ionic Solute (charges)	INSOLUBLE	SOLUBLE

Factor 2 – *Temperature*

- For solids, as temperature ↑, solubility ↑

  Remember: Table G is there to help you!
  Gases slope ↓ solids slope ↑.

- For MOST gases, as temperature ↑, solubility ↓
  Example: heat soda → carbonation is LOST!
  (warm)

Factor 3 – *Pressure*

- For solid or liquid solutes → LITTLE TO NO EFFECT
- For gases dissolved in liquids → As pressure increases, VOLUME ↓, SOLUBILITY ↑

  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 

  When packaged, soda is pressurized (so CO_2 dissolves in the liquid). When you open the can, pressure is let out (pressure decreases in the can). So the gas particles are released and you get CARBONATION!

$AB + CD \rightarrow AC + BD$

SOLUBILITY - TABLE F:

- Tells you what substances are soluble or insoluble
- Is used to predict whether or not a precipitate will form when two ionic solutions are mixed and undergo a double replacement reaction

  Precipitate = INSOLUBLE product that falls to the bottom of a mixture/solution (following a double replacement rxn.)

- A precipitate will form (or a visible reaction will occur) if one or both of the products listed is insoluble (this is a precipitate!)

Table F
Solubility Guidelines for Aqueous Solutions

Ions That Form *Soluble* Compounds	Exceptions (insoluble)
Group 1 ions (Li^+, Na^+, etc.)	
ammonium (NH_4^+)	
nitrate (NO_3^-)	
acetate ($C_2H_3O_2^-$ or CH_3COO^-)	
hydrogen carbonate (HCO_3^-)	
chlorate (ClO_3^-)	
perchlorate (ClO_4^-)	
halides (Cl^-, Br^-, I^-)	when combined with Ag^+, Pb^{2+}, and Hg_2^{2+}
sulfates (SO_4^{2-})	when combined with Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, and Pb^{2+}

Ions That Form *Insoluble* Compounds	Exceptions (soluble)
carbonate (CO_3^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
chromate (CrO_4^{2-})	when combined with Group 1 ions, Ca^{2+}, Mg^{2+}, or ammonium (NH_4^+)
phosphate (PO_4^{3-})	when combined with Group 1 ions or ammonium (NH_4^+)
sulfide (S^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
hydroxide (OH^-)	when combined with Group 1 ions, Ca^{2+}, Ba^{2+}, Sr^{2+}, or ammonium (NH_4^+)

COLUMN 1 COLUMN 2 COLUMN 3 COLUMN 4

HOW TO USE TABLE F:

- **Find one of the ions** in your formula in either **Column 1 or Column 3**
- Check if there are any exceptions to that ion (glance over to Column 2/Column4)
- If there are no exceptions, scroll up Column 1 or 3 (whichever you are in) to find if your compound is soluble or insoluble
- If there is an exception (that ion listed in column 2 or 4 as an exception will be what your ion is bonded to), then your answer is the heading to Column2 or 4.

PRACTICE: Using Table F, determine whether the following are **SOLUBLE** or **INSOLUBLE**

1) **KNO_3** Soluble
2) **$CaCO_3$** Insoluble
3) **$MgSO_4$** Soluble
4) **NH_4OH** Soluble

***<u>SOLUBILITY</u>:** a measure of how much SOLUTE will dissolve in a certain amt. of SOLVENT @ a given temp.

- **Soluble:** will dissolve; AKA miscible

  **The more soluble something is, the more concentrated it can be in soln*

- **Insoluble:** will NOT dissolve

- **Concentration:** RATIO of the amt. of solute per amount of solvent (SOLUTE/SOLVENT)

<u>DILUTE</u>: soln that has a relatively SMALL amount of SOLUTE dissolved into a relatively LARGE amount of SOLVENT (WEAK solution)
(verb: to WEAKEN the CONCENTRATION of a soln)

$$\text{Dilute} = \frac{\text{solute}}{\text{SOLVENT}}$$

<u>CONCENTRATED</u>: Solution that has a relatively LARGE amount of SOLUTE dissolved into a relatively SMALL amount of SOLVENT (STRONG solution)

$$\text{Concentrated} = \frac{\text{SOLUTE}}{\text{solvent}}$$

Most vs. Least Concentrated (Using Table F)

(STRONG) (WEAK)

- CONCENTRATION is directly related to amt. of solute DISSOLVED
  - REGENTS QUESTION:
    - Most concentrated = SOLUBLE compounds
    - Least concentrated = INSOLUBLE compounds

**If a compound doesn't dissolve, it doesn't add to the concentration of the soln*

REACTIONS AND TABLE F:

Click here to watch vodcast: https://www.youtube.com/watch?v=tB4O09KnFVo&feature=youtu.be

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

$$2AgNO_3 + Na_2CrO_4 \rightarrow Ag_2CrO_4 + 2NaNO_3$$

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

$$Ag_2CrO_4 = \text{ppt}$$

NOW YOU TRY SOME:

1) Will a precipitate form when lithium nitrate is mixed with potassium phosphate?

$$3\overset{+}{Li}\overset{-}{NO_3} + \overset{+}{K_3}\overset{-3}{PO_4} \rightarrow Li_3PO_4 + 3KNO_3 \quad = \quad \text{no ppt}$$

For the following reactions, determine the products formed and then circle any precipitate formed as a product:

2) $\overset{+}{Ag}\overset{-}{NO_3}$ + $\overset{+}{K}\overset{-}{Cl}$ → (AgCl) + KNO_3

3) $\overset{+2}{Cu}\overset{-2}{SO_4}$ + $\overset{+}{Na}\overset{-}{OH}$ → ($\overset{+2}{Cu}(\overset{-1}{OH})_2$) + $\overset{+1}{Na_2}\overset{-2}{SO_4}$

4) Li_2CrO_4 + NH_4NO_3 → $LiNO_3$ + $(NH_4)_2CrO_4$ No ppt

Calculating [Concentration] - all these formulas can be found on Table T!

Use the following methods/formulas to calculate the concentrations of the solutions below. SHOW YOUR WORK.

Molarity (M)

Formula:

molarity = moles of solute / liters of solution

Example: What is the molarity of a solution that contains 4.0 moles of NaOH in 0.50 L solution?

molarity = 4.0 mol / 0.5 L = 8.0 M

Parts per Million (ppm)

Formula:

PPM = mass of solute / mass of whole × 1,000,000

Example: Approximately 0.0043 g of oxygen can be dissolved in 100. mL of water at 20°C. Express this in terms of part per million.

= 100 g H_2O

PPM = 0.0043 g / 100 g × 1,000,000 = 43 PPM

Percent by Mass

Formula: % Comp = mass of part / mass of whole × 100

Example: What is the percent by mass of sodium hydroxide if 2.50 g of NaOH is added to 50.00 g of H_2O?

% Comp = 2.50 / 50.0 × 100 = 4.76%

Percent by Volume

Formula: % Comp = mass of part / mass of whole × 100

Example: What is the percent by volume of alcohol if 50.0 mL of ethanol is diluted with water to form a total volume of 300. mL?

% Comp = 50 mL / 300 mL × 100 = 16.7%

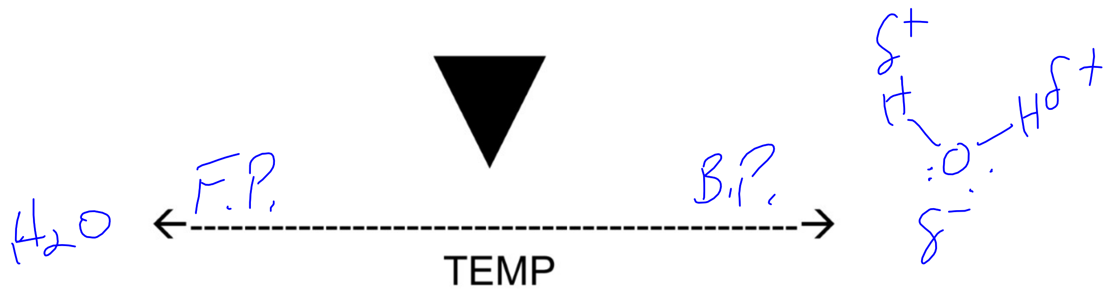
COLLIGATIVE PROPERTIES FOR ELECTROLYTES

(Based on NUMBER OF PARTICLES)

Electrolyte = A substance that produces IONS when dissolved in a solution. Because the ions are FREE TO MOVE AROUND (MOBILE) in the solution, the solution is able to CONDUCT electricity (salts).

*Same ***concept*** as mobile electrons in metals, but not the same thing. Don't get confused on the Regents!!!

When a solute is dissolved in a solvent, solvent molecules surround the particles of the solute. This causes the boiling point and freezing point of the solution to change in a very specific and predictable way.



Boiling Point Elevation = b.p. INCREASES when solute is added

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

Freezing Point Depression = f.p. DECREASES when solute is added

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

IONIC VS MOLECULAR SOLUTES:

* Why do we salt the roads in the winter rather than sugar them?!?!?

	MOLECULAR: $C_{12}H_{22}O_{11}$ (sucrose – covalent!)	IONIC: Salt (NaCl)
Reaction produced when solute dissolved in water	$C_{12}H_{22}O_{11(s)} \xrightarrow{H_2O} C_{12}H_{22}O_{11(aq)}$ 1 mol 1 mol	$NaCl_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$ 1 mol 2 mol
Number of moles of product produced	1 mol → 1 mol	1 mol → 2 mol
General Rule Example: $CaCl_2$ # ions produced:	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) $CaCl_{2(s)} \xrightarrow{H_2O} Ca^{2+}_{(aq)} + 2Cl^-_{(aq)}$ 1 mol → 3 mol	

**$CaCl_2$ i$ even better than NaCl, $o why don't we u$e it to $alt our road$?*

Becau$e IT COSTS MORE MONEY!

Freezing point depressions (laboratory conditions):

NaCl = -29.4ºC (-21ºF) CaCl₂ = -51ºC (-60ºF)

2 mol 3 mol

**NaCl becomes ineffective as a melting agent at around -9ºC (15ºF). $CaCl_2$ works effectively at temperatures as low as -29ºC (-20ºF).