# Solubility and Solutions

Solutions can exist for any combination of solids, liquids and gases.
In this chemistry class we will primarily focus on liquid water solutions (aqueous).

**Solutions are homogeneous mixtures** of very small particles.

Mixtures are physical combinations. Parts of mixtures can be separated by **physical methods**. However; the particles in a solution are so small that they **can** pass through filter paper. A filter cannot separate parts of a solution.
Some methods to separate components (parts) of solutions are **chromatography**, **distillation**, and **evaporation**. (Think back to the topic of "Matter").
Solutions are transparent. You can see through a solution. Even if solutions are colorful, such as solutions of transition metal compounds, you can still see through them.

The 2 parts of a **solution** are **solute** and **solvent.**

**What are the 2 parts of a solution?** ____________________ and ________________.

**Solute** refers to the substance that dissolves.
**Solvent** refers to the substance that the solute dissolves into.

The classic example is a **saltwater solution.** The ***salt*** is the ***solute*** and the **water** is the **solvent**. Together they form a **solution**. Remember, a water solution is referred to as **aqueous**. The abbreviation for aqueous is (aq). Therefore; $NaCl_{(aq)}$ describes a solution with water as the solvent and NaCl as the solute. Since all solutions are homogeneous, $NaCl_{(aq)}$ , is a homogeneous solution.

**Describe and provide an example for:**

Solute

_______________________________________________________________

Solvent

_______________________________________________________________

**Solubility** describes the "dissolvability" of a substance. If something is soluble, it is able to dissolve. The greater a substance's solubility, the more readily (easily) it dissolves. If a substance does not dissolve, it is said to be **insoluble**. You may recall the rule for solubility, **"like dissolves like"**. Polar substances dissolve in other polar substances. Don't forget water is polar (polar bears like water). Non-polar substances dissolve in other non-polar substances. Oil is the classic example of a nonpolar liquid.

**Describe:** solubility ______________________________________________

**Immiscible** refers to liquids that are not soluble in each other. Oil and water do not dissolve in each other because of different polarities and are called **immiscible.** Such a combination is not homogeneous and therefore not a solution. You can think of immiscible as **not mix able**.

**Describe:** immiscible ______________________________

**Saturated** describes a solution that is holding as much solute as possible for a given temperature. If you add a small amount of sugar to water it will dissolve. If you keep adding sugar, eventually it will stop dissolving. Why? The solution is full. It can hold no more. We refer to this solution as **saturated.**

**Describe:** saturated ______________________________

**Unsaturated** describes a solution can have more solute dissolved into it. Think about when you are making a glass of your favorite iced tea from a powdered mix. When you add one scoop of mix, the mix dissolves and the drink can still have more iced tea mix dissolved into it. As you keep adding mix you will eventually reach a point when no more will dissolve and the powder will just sit at the bottom undissolved. When no more powder can be dissolved, the solution is saturated.

**Describe:** unsaturated______________________________

**Supersaturated** refers to a solution that holds more solute than normal solubility rules will allow.
A **supersaturated solution** can be prepared by adding solid solute to a hot solvent until it is saturated. The solution is then slowly cooled. This results in more solute being dissolved than would normally be for a given temperature. A supersaturated solution is unstable and the solute may rapidly leave the solution all at once.

**Describe:** supersaturated ______________________________

**Precipitate** describes a solid that forms from a solution. As a saturated solution is cooled, some of its solute will leave the solution. The resulting solid is called a precipitate.

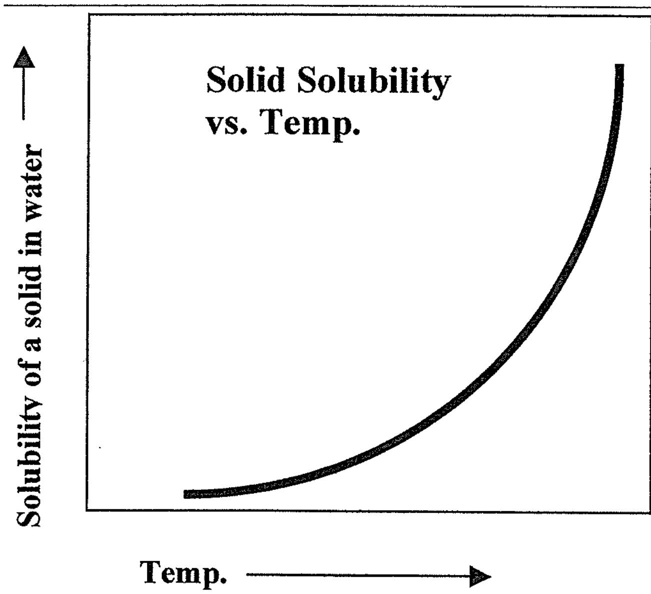
**Describe:** precipitate ______________________________

# Solubility of solids in liquids

**Rule: solubility of a solid increases as temperature (kinetic energy) increases.** This should seem familiar. As you probably know, you can dissolve more sugar in hot tea than in cold tea. What's that you say? You like coffee instead of tea. Don't worry it still works the same way. 

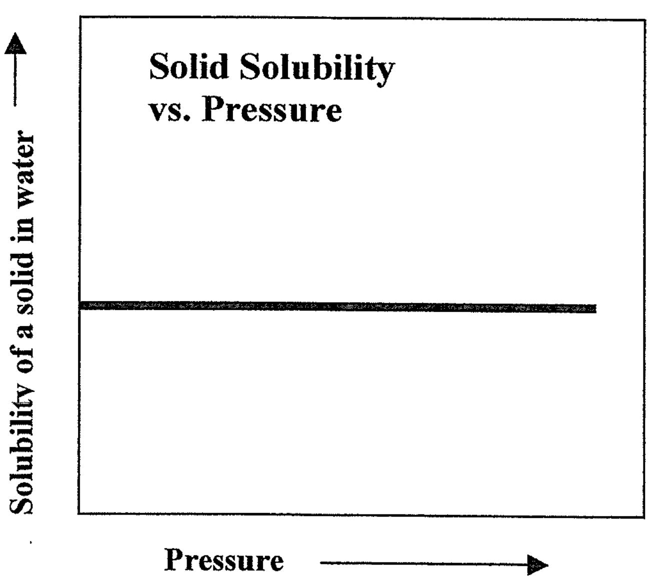
**What is the rule for solubility of solids in liquids?** ______________________________

________________________________________________________________________




Notice, the line goes up. An increase in temperature increases the solubility of a solid in a liquid

**Pressure** has **no effect on the solubility of solids** in liquids.




Notice, the line is flat. A change in pressure has no measurable effect on solid solubility.

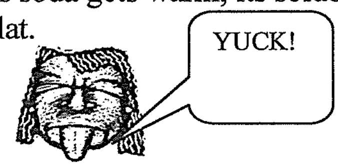
# Solubility of Gases in Liquids

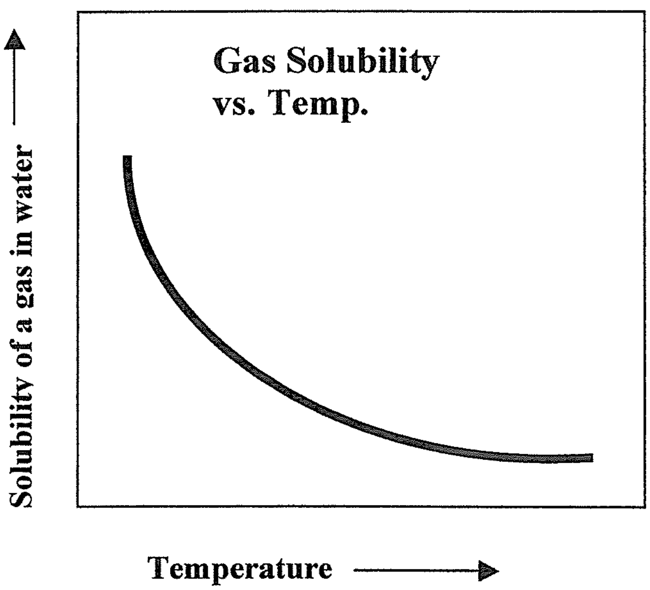
There are 2 factors for solubility of gases in liquids.
**Temperature** and **Pressure**

**What are the 2 factors for the solubility of gases in liquids?**

______________________________ and ________________________________

**Solubility of a gas in a liquid decreases as temperature (kinetic energy) increases. Why?** Gas particles have more kinetic energy at higher temperatures and can escape from the solvent meaning less gas dissolved in solution.
Where have you observed this in your daily lives? Think of the fizzy bubbles in soda. The fizz in soda results from dissolved carbon dioxide gas ($CO_2$) escaping from the liquid. When soda is cold it holds a lot of $CO_2$. As soda gets warm, its solubility decreases and the $CO_2$ escapes leaving the soda flat.




Notice, the line goes down. An increase in temperature decreases the solubility of a gas in a liquid. A cold liquid holds more dissolved gas than a warm liquid.

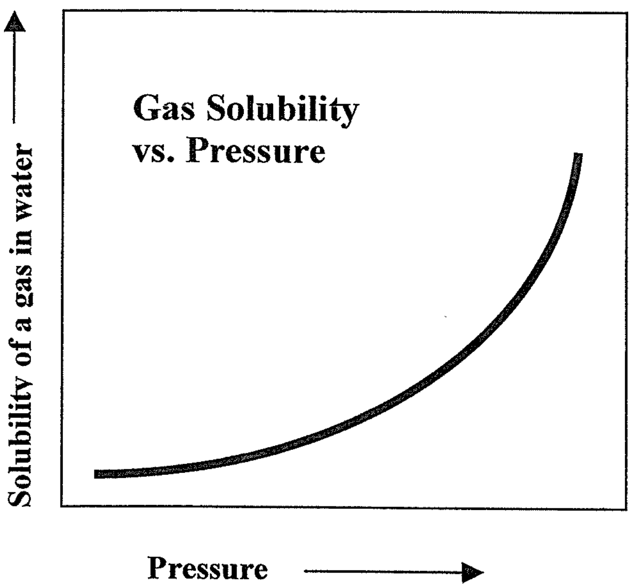
**Explain the relationship between gas solubility and temperature.** ______________

______________________________________________________________

**Gas solubility in a liquid increases as the pressure increases.**

**Why?** At higher pressures more gas particles strike the surface of the solution and enter the solvent, making the concentration of the gas higher.

We are reminded of this again when we think of soda. How would you describe the pressure inside a new unopened bottle of soda? That's right, Very high! Once the bottle is open the pressure decreases and the soda starts to go flat as the $CO_2$ leaves the solution.

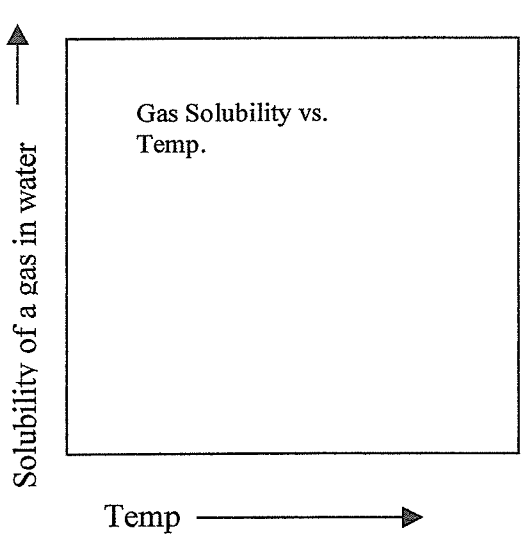



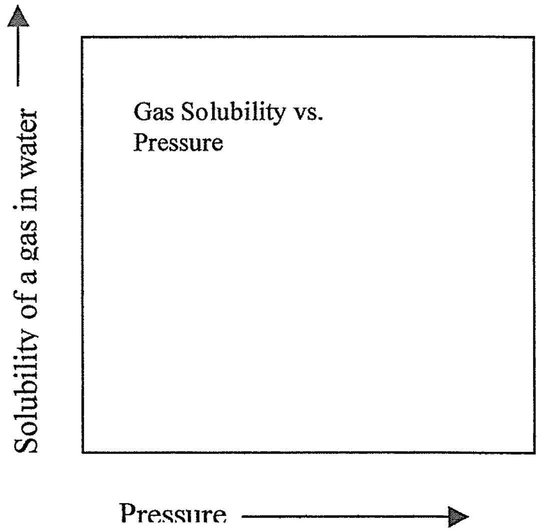

Notice, the line goes up. An increase in pressure increases the solubility of a gas in a liquid

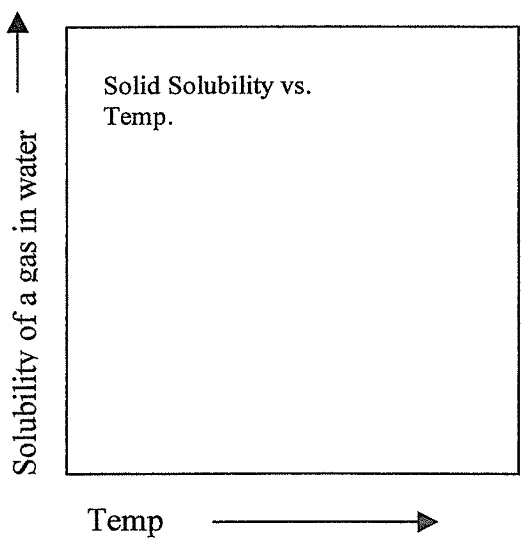
**Explain the relationship between gas solubility and pressure.** ______________

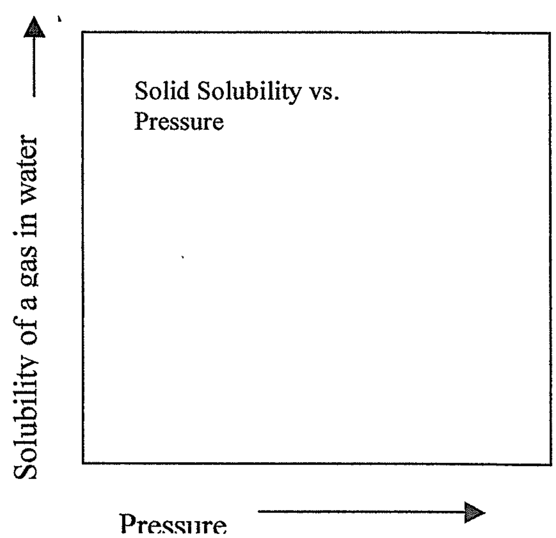
______________________________________________

**Sketch lines to represent the following graphical relationships.**



# Table G Solubility Curves

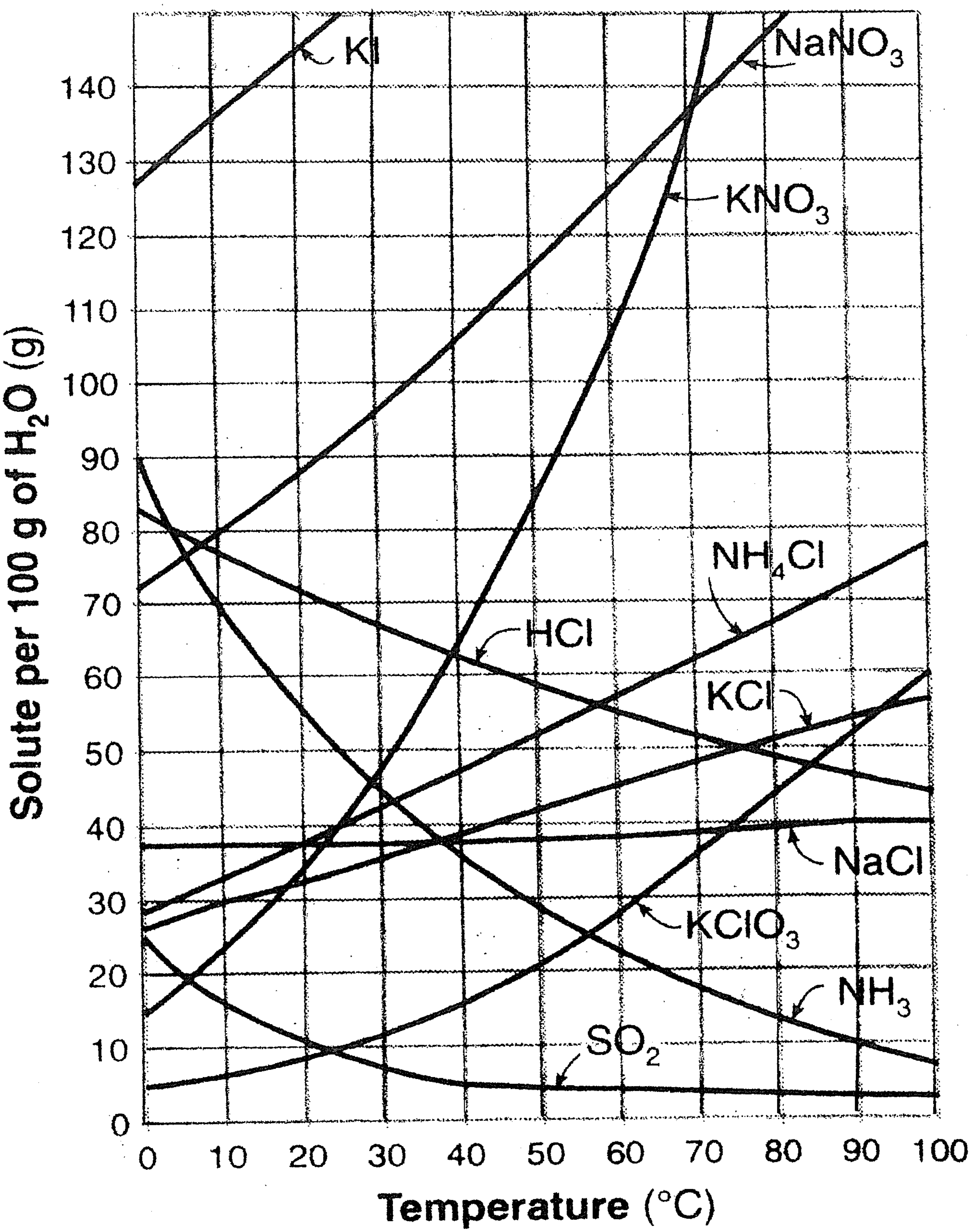
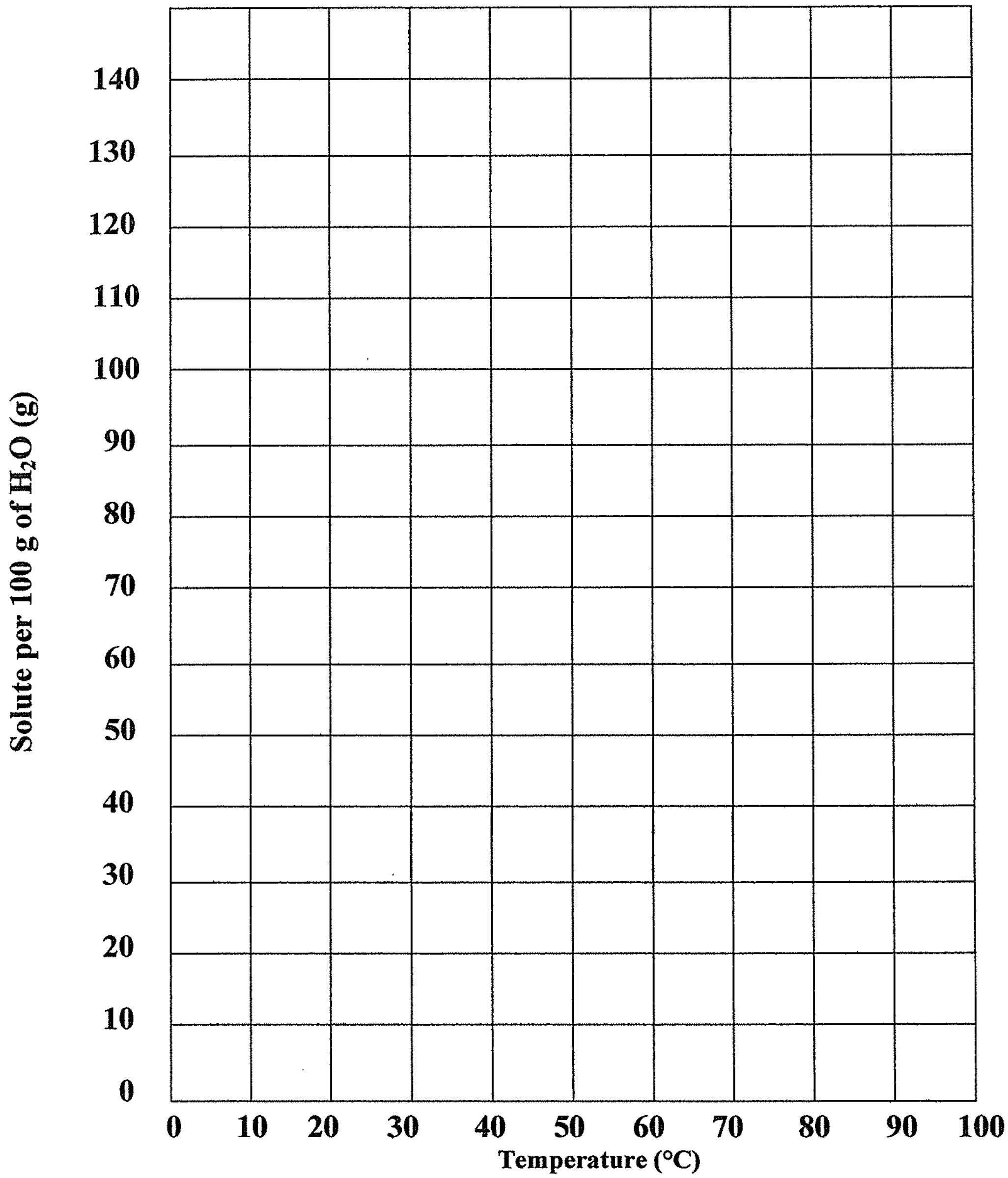


Table G




**Let's try to understand what all the lines on Table G are showing us.**

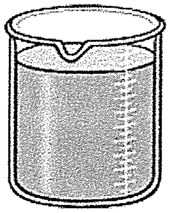
There are many lines, at first, it looks very confusing. ☹

Plot the following data collected from an experiment to check the solubility (dissolvability) of potassium nitrate ($KNO_3$) in 100 g (100 mL) of water at a variety of temperatures. ✎

| Mass of $KNO_3$ | 15 g | 23 g | 34 g | 47 g | 64 g | 84 g | 106 g | 137 g |
|---|---|---|---|---|---|---|---|---|
| Water Temp °C | 0 °C | 10 °C | 20 °C | 30 °C | 40 °C | 50 °C | 60 °C | 70 °C |

**The resulting line is a solubility curve. It shows us how much solute can be dissolved in 100 g of water at many temperatures. All the rest of the lines on Table G show the same information for other substances.**

**Table "G"** is used to determine the solubility of various solids and gases in a water solution. Table "G" assumes **100 grams of water** as its standard. Remember.




**100 mL** of liquid water has a mass of **100 grams**! How do you interpret all the lines and numbers on Table "G"? Notice that the X-axis (the bottom) is temperature in °C. Notice that the Y-axis (the left side) is grams of solute per **100 g** of water. Think of this table as referring to a beaker containing **100 mL** of water. All the lines that increase with temperature are for solids. All the lines that decrease with temperature are for gases. Hopefully, you remember that $NH_3$ is ammonia. You might think that ammonia is a liquid because there is a bottle labeled "ammonia" under your kitchen sink that contains a liquid. However; the household cleaner, ammonia, is actually $NH_{3(aq)}$. That describes ammonia gas dissolved in water.

The curved lines on Table"G" represent the points at which **100 g** of water will be saturated with a particular substance at any given temperature. Let's see how this works!

**Ex: How many grams of ammonium chloride ($NH_4Cl$) are required to saturate 100 g of water at 50° C?**

To solve, find 90° C on Table "G" and follow it up until it intersects the curved line for $NH_4Cl$. Once you reach the intersection point, follow the line to the left until you reach the Y-axis. This is where you read the # of grams.

Answer: approximately **52 g.**

**Ex: How many grams of ammonium chloride ($NH_4Cl$) are required to saturate 100 g of water at 90° C?**

Answer: approximately **72g.**

Notice from the 2 above examples, the solubility of ammonium chloride (a solid) increased as temperature increased.

Now you try it.

How many grams of potassium chlorate ($KClO_3$) are required to saturate 100 g of water at 50° C?

As mentioned earlier, the solubility of a gas in a liquid decreases as temperature increases. Let's examine ammonia gas ($NH_{3(g)}$).

**Ex: How many grams of ammonia ($NH_{3(g)}$) are required to saturate 100 g of water at 50° C?**

Answer: approximately **28 g.**

**Ex: How many grams of ammonia ($NH_3$) are required to saturate 100 g of water at 90° C?**

Answer: approximately **10 g.**

Notice from the 2 gas examples, the solubility of ammonia (a gas) decreased as temperature increased.

**Important!** When you use Table G, the amount of water (solvent) is 100 g. The table determines the amount of solute that can be dissolved in 100 g of water (solvent) for a range of temperatures.

If a question uses an amount of water other than 100 g, you must make the proper adjustment.

**Ex: How many grams of potassium nitrate ($KNO_3$) are required to saturate 200 g of water at 60°C?**

To solve this problem locate the 60°C line. Follow the line upward until it intersects the curved line for a potassium nitrate ($KNO_3$). From this intersection point, follow it to the left and read the # of grams. The amount is **105 g**. However; you must remember, that 105 grams is for 100g of water and the question is concerned with 200 g of water. What do you have to do? You guessed it. Since there is twice as much water, double the amount required for 100 g. The final answer is **105 g X 2 = 210 g.**

Now you try it.

**a) How many grams of potassium chloride (KCl) are required to saturate 300 g of water at 10°C?**

**b) ) How many grams of potassium chloride (KCl) are required to saturate 250 g of water at 80°C?**

**c) ) How many grams of potassium nitrate ($KNO_3$) are required to saturate 50 g of water at 50°C?**

## Points of Intersection on Table G

A given temperature (°C) and amount of solute (grams/100 g $H_2O$) provide a point of intersection on Table G.
You will have to provide information for points of intersection on Table G.
You will have to know if an intersection point represents a **saturated** solution, an **unsaturated** solution, or a **supersaturated** solution. To answer this type of question, first you must find the intersection point. Next you will have to determine if the point is on the line for the substance, above the line, or below the line. If the point is exactly on the line the solution is saturated. If the point is above the line the solution is supersaturated. If the point is under the line the point is unsaturated.
You may find it helpful to remember it this way:

- If the point **sat** on the line, the solution is **sat**urated.
- If the point is **un**der the line, the solution is **un**saturated.
- If the point is above the line, like **super**man flying above Metropolis, the solution is

**Super**saturated.

Don't forget. When using an amount of water other than 100 g you must adjust the amount of **solute** correspondingly.
If there is twice as much water (200 g) **you must divide it by 2** to change it to 100 g so you can use Table G. Next, **you must also divide the amount of solute by 2.**
If there is half as much water (50 g) **you must multiply it by 2** to get to 100 g so you can use table G. Next, **you must multiply the amount of solute by 2.**

**Examples: Use the terms saturated, unsaturated, or supersaturated to describe the following solutions.**

A) 100 g of 50°C water containing 30g of NaCl.

B) 100 g of 90°C water containing 10g of $NH_3$.

C) 100 g of 50°C water containing 130g of $NaNO_3$.

D) 100 g of 50°C water containing 10g of $KClO_3$.

E) 100 g of 90°C water containing 80g of NaCl.

F) 200 g of 10°C water containing 160g of $NaNO_3$.

G) 50 g of 70°C water containing 60g of $KNO_3$.

H) 50 g of 90°C water containing 30g of $NH_4Cl$.

## More difficult problem.

**As a saturated solution (solid solute) cools, some solute will precipitate (leave the solution as a solid).** When this happens, some dissolved solute will **precipitate** (leave the solution and usually sink to the bottom).

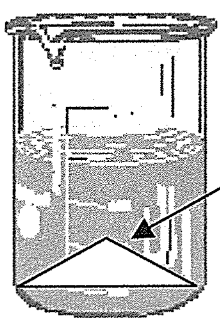
You will need to look at Table G to better understand this.

**Ex: 100 g of water contains 70 g of $NH_4Cl$ at 85 °C. If the solution cools to 10 °C, how much $NH_4Cl$ will precipitate (settle to the bottom)?**

Step 1: Locate the intersection point for **70 g** of $NH_4Cl$ at 85 °C, it is on the line for $NH_4Cl$.
The point **sat** on the line. This solution is saturated.

Step 2: Determine the # of grams that the solution will hold at 10 °C. answer **35 g.**

Step 3: Subtract 35 g from 70 g. 70 g – 35 g = **35 g of $NH_4Cl$ will precipitate.**

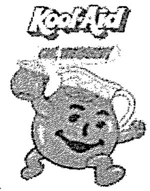


Try one on your own.

**100 g of water contains 54 g of KCl at 90 °C. If the solution cools to 60 °C, how much KCl will precipitate (settle to the bottom)?**

# Solution Concentration

The term **concentration** describes how much solute is in a solution. If you add 1 scoop of




Kool Aid to a pitcher of water, the resulting solution will have a low concentration. Solutions with a low concentration are called **dilute.**
If you add 20 scoops of Kool Aid to a pitcher of water the resulting solution will have a high concentration. Solutions with a high concentration are simply called **concentrated.**

## Measuring Concentration

Solution concentration can be expressed n terms of:
**molarity "M", parts per million "ppm", percent by mass, or percent by volume.**
The formulas used to calculate molarity and ppm are available in "Table T" pg. 12 of the reference table. **Percent by mass and percent by volume are normal percentage calculations.**
The symbol for molarity is **M.**

The formula to calculate molarity is. $$M = \frac{\text{moles of solute}}{\text{Liters of solution}}$$

**Remember!** The molarity formula uses liters. If the given information in the question uses milliliters, you must convert milliliters to liters. Do you remember how to do that? Of Course! , But I'll remind you anyway. **There are 1000 mL in a liter.**
Therefore; 500 mL = .5 liter. 250 mL = .25 L. 832 mL = .832 L.
To convert from milliliters to liters is to move the decimal 3 places to the left.

You try it! How many liters are = to 750 mL? _______________ **Good work!**

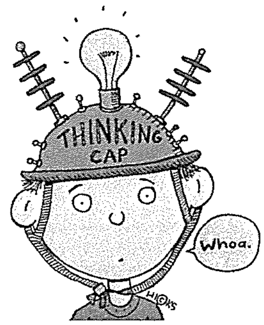
**Ex: What is the molarity of a solution if 3 moles of NaOH are dissolved in 4.0 liters of water?**

$$M = \frac{\text{Moles of solute}}{\text{liters of solution}} \qquad M = \frac{3}{4} \qquad M = .75$$

**Ex:** What is the molarity of a solution if 2 moles of NaCl are dissolved in 3 liters of water? **Show Work!**

Just like any formula with 3 variables, you may be given any 2 things in the molarity formula and have to solve for the unknown.

**Ex:** How many moles of LiCl (solute) are required to make 1.5 liters of a 5M solution? **Show Work!**




**Put on your thinking cap!**

Some problems may require you to combine the mole calculation formula and the molarity formula. Don't forget you can find the formula for mole calculations on pg. 12 "Table T" of the reference table.

$$\text{\# moles} = \frac{\text{given mass}}{\text{gfm}}$$

**Ex: What is the molarity of a solution if 80 g of NaOH are dissolved in 800 mL of water?**

Step # 1 Determine how many moles of NaOH are present in 80g.

$$\text{\# moles} = \frac{\text{given mass}}{\text{gfm}} = \frac{80\text{ g}}{40\text{g/mol}} = \textbf{2}\text{ moles NaOH present.}$$

Step # 2 Determine the # of liters present **800 mL = .8 L**

Step # 3 Substitute into the molarity formula and solve.

$$\mathbf{M} = \frac{moles..solute}{Liters..solution} = \frac{2mol}{.8L} \qquad \mathbf{M = 2.5}$$

Now you try some!

a) What is the molarity of a solution if 116 g of NaCl are dissolved in 900 mL of water?

b) How many grams of lithium chloride are contained in .250 L of a 1.75 M solution?

# Parts per Million

The formula to calculate parts per million can be found on pg. 12 "Table T" of the reference table.

$$\textbf{Parts per million} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 1\,000\,000$$

**Ex: Describe the concentration of a solution in parts per million (ppm) if the solution contains .002 g of lead in 10 000 g of water solution.**
To solve, insert the given data into the formula.

$$\text{Parts per million} = \frac{.002\text{ g}}{10\,000\text{ g}} \times 1\,000\,000 \qquad = \mathbf{.2\ ppm}$$

**Practice Problems:**

a) Describe the concentration in ppm if the solution contains .03g of arsenic in 20 000 g of water solution.

b) Express the concentration in ppm if .25 g of sugar is dissolved in 750 g of solution.

## Solution Concentration as **Percent by Volume.**

Calculating any percentage uses the same concept and formula.

**Percentage =** $\frac{part}{whole}$ **X 100%**

**The volume of solute is the part and the volume of solution is the whole.**

You must remember that the solution (*whole*) is the solute + solvent.

**Ex: What is the percent by volume of a solution made by dissolving 20 mL of alcohol in 50 mL of water?**

Step 1: Find the total volume of solution by adding the volume of solute and solvent

20 mL alcohlol (solute) + 50 mLwater (solvent) = 70 mL (solution).

Step 2: Insert the information into the percentage formula.

**Percentage =** $\frac{(solute.volume)part}{(solution.volume)whole}$ **X 100%**

**Percentage by volume=** $\frac{20mL}{70mL}$ **X 100%**

**Percentage by volume = .29 X 100% = 29%.**

**Practice percent by volume problems.**

1) What is the percent by volume of 100 mL of alcohol dissolved in 800 mL of water?

2) What is the percent by volume of 150 mL of acetone dissolved in 1800 mL of alcohol?

## Solution Concentration as **Percent by Mass.**

Calculating any percentage uses the same concept and formula.

**Percentage** = $\frac{part}{whole}$ **X 100%**

**The mass of solute is the part and the mass of solution is the whole.**

You must remember that the solution (*whole*) is the solute + solvent.

**Ex: What is the percent by mass of a solution made by dissolving 90.0 g of calcium chloride in 950.0 g of water?**

Step 1: Find the total mass of solution by adding the mass of solute and solvent

90.0 g calcium chloride (solute) + 950.0 g water (solvent) = 1040.0 g (solution).

Step 2: Insert the information into the percentage formula.

**Percentage** = $\frac{(solute.mass)part}{(solution.mass)whole}$ **X 100%**

**Percentage by mass**= $\frac{90.0g}{1040.0g}$ **X 100% Percentage by mass = .087 X 100% = 8.65%**

**Practice percent by mass problems.**

1) What is the percent by mass of 100.0 g of salt dissolved in 2000.0 g of water?

2) What is the percent by mass of 15.0 g of iodine dissolved in 100.0 g of alcohol?

# Changing Solution Concentration

The following formula is used when the volume and concentration of a solution change due to adding solvent (**dilution**) or removing solvent (concentration by **evaporation**).

The formula is $M_1 \times V_1 = M_2 \times V_2$ This formula is not available in the reference table but looks very much like the **titration** formula. $M_A \times V_A = M_B \times V_B$

$M_1$ = initial molarity
$V_1$ = initial volume
$M_2$ = final molarity
$V_2$ = final volume

**Example:** Let's say we had a 1000.0 mL beaker containing 300.0 mL of 4.0 M $HCl_{(aq)}$
If we add 200.0 mL of water to the solution, what will be the new concentration (molarity)?
Logically, the new concentration (molarity) will be less than the original, since the amount of solute is the same but there is now more solution.
To solve:
Step 1: Choose the correct equation $M_1 \times V_1 = M_2 \times V_2$

Step 2: List the known information
$M_1$ = initial molarity = 4.0 M
$V_1$ = initial volume = 300.0 mL
$M_2$ = final molarity ?
$V_2$ = final volume 300.0 mL + 200.0 mL = 500.0 mL

Step 3: Substitute into the equation $4.0\text{ M} \times 300.0\text{ mL} = M_2 \times 500.0\text{ mL}$

Step 4: Solve $\frac{4.0\text{ M} \times 300.0\text{ mL}}{500.0\text{ mL}} = M_2$

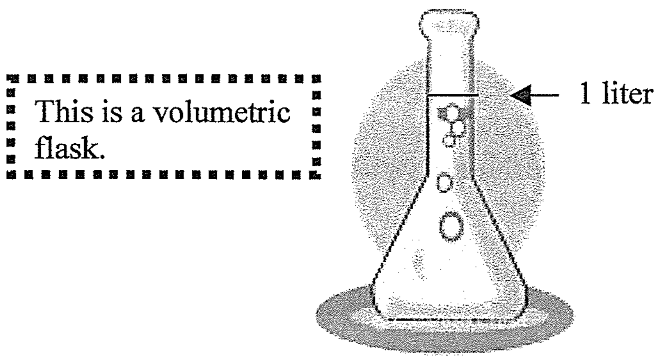
$M_2 = 2.4\text{ M}$

Try this **example: Show all Work**
If a 500 mL beaker containing 400.0 mL of 1.0 M $CuSO_{4(aq)}$ is allowed to evaporate until the new volume is 350.0 mL, what is the new concentration (molarity)?

# How to Prepare a Standard Solution from Solid Solute.

Here are the steps involved in the preparation of a standard solution from a solid solute.

Step 1: A weighed amount of solute is put into a volumetric flask and small quantity of water is added.




Step 2: The solid is dissolved in the water by gently swirling the water in the flask.

Step 3: More water is added, until the level reaches the measured mark on the flask.

**Describe how to prepare 1 liter of 2.00 M $KNO_3$ solution, using a 1 liter volumetric flask.**

## How to Prepare a Solution of a Specific Molarity from a Standard Solution.

A *standard* solution is a concentrated solution with a known molarity. A *standard* solution can be used to prepare a solution of desired molarity.

Describe standard solution ______________________________________________

We use the formula: $M_1 \times V_1 = M_2 \times V_2$

$M_1$ = initial molarity $V_1$ = initial volume $M_2$ = final molarity $V_2$ = final volume

***WARNING!***

When water is added to a strong acid it is a very exothermic process. If water is added to acid, the water may boil and splatter hot acid .The proper way to prepare acid solutions is to slowly add the acid to the water while stirring. **Remember! AA** for Add Acid.

**Ex: How much 12.0 M $HCl_{(aq)}$ (hydrochloric acid solution) is required to prepare 500.0 mL of 3.0 M $HCl_{(aq)}$ (hydrochloric acid solution)?**
**Describe how to prepare the solution.**
To solve:
Step 1: Choose the correct equation $M_1 \times V_1 = M_2 \times V_2$

Step 2: List the known information
$M_1$ = initial molarity = 12.0 M
$V_1$ = initial volume = ? mL
$M_2$ = final (desired) molarity = 3 M
$V_2$ = final (desired) volume = 500.0 mL

Step 3: Substitute into the equation $12.0\ M \times V_1 = 3.0\ M \times 500.0\ mL$

Step 4: Solve $V_1 = \frac{3.0\ M \times 500.0\ mL}{12.0\ M}$

$V_1$ = 125.0 mL

If you want 500 mL of 3 M $HCl_{(aq)}$ you will have to add the 125.0 mL of 12.0 M $HCl_{(aq)}$ to 375.0 mL of water. 125.0 mL + 375.0 mL = 500.0 mL.

To make 500 mL of 3 M $HCl_{(aq)}$ slowly add the 125.0 mL of 12.0 M $HCl_{(aq)}$ to 375.0 mL of water while stirring. Don't forget to wear goggles!

Try one on your own.
How much 18 M $H_2SO_{4(aq)}$ (sulfuric acid solution) would be required to prepare 300 mL of 6M $H_2SO_{4(aq)}$ (sulfuric acid solution).
Describe how to prepare the solution.

# Understanding Table "F"

## Solubility Guidelines for aqueous Solutions

Table "F" determines the solubility of ionic compounds in water. **Solubility** means **dissolvability**. If a substance is soluble in water it dissolves in water. If a substance is **insoluble** in water it does not dissolve in water. Many ionic compounds are soluble in water, but not all of them. Ionic compounds are composed of positive and negative ions. The ions can be ions of simple elements ex: $Ba^{+2}$, $Na^{+}$, $Sr^{+2}$, $S^{-2}$, etc. The ions may be polyatomic ions ex: $NH^{+4}$, $NO_3^{-}$, $CO_3^{-2}$, etc.
Remember! Names, formulas, and charges of polyatomic ions can be found in Table "E".

**Table "F" consists of two tables.**

The **left table** identifies ions that form **soluble** compounds.
The **right table** identifies ions that form **insoluble** compounds.

If a compound contains an ion from the **left** column of the **left** table it will be soluble unless it is combined with an ion from its exceptions column.
Ex: Any compound that contains a Group 1 ion or ammonium ($NH_4^{+}$) is soluble.

Ex: Any compound that contains a halide (Group 17) ion ($F^{-}$, $Cl^{-}$, $Br^{-}$, or $I^{-}$) will be soluble unless combined with $Ag^{+}$, $Pb^{+2}$, or $Hg_2^{+2}$.

If a compound contains an ion from the **left** side of the **right** table it will be **insoluble** unless it is combined with an ion from the exceptions column.
Ex: Any compound that contains phosphate ($PO_4^{-3}$) is insoluble unless it is combined with a Group 1 ion or ammonium ($NH_4^{+}$).

Ex: Any compound that contains chromate ($CrO_4^{-2}$) is insoluble unless it is combined with a Group 1 ion, $Ca^{+2}$, $Mg^{+2}$ or ammonium ($NH_4^{+}$).

**Sample Problem:** Determine if the following compounds are soluble.
a) aluminum sulfate $Al(NO_3)_3$
Notice that this compound is made of aluminum and nitrate.
**Step 1:** Look for aluminum or nitrate in the left column of either table. Aluminum is not listed but nitrate is in the left column of the soluble table.
**Step 2:** Check for exceptions. There are no exceptions for nitrate. All nitrate compounds are soluble. Therefore, aluminum sulfate "$Al(NO_3)_3$"is soluble.

b) copper II carbonate "$CuCO_3$"
Notice that this compound if made of copper and carbonate.
**Step 1:** Look for copper or carbonate in the left column of either table. Copper is not listed but carbonate is in the left column of the insoluble table.
**Step 2:** Check for exceptions. There are exceptions for carbonate but copper is not one of them. Therefore, copper II carbonate "$CuCO_3$" is insoluble.

**Important! Only soluble items can form a solution. Insoluble items will not form a solution.**

# Table F
# Solubility Guidelines for Aqueous Solutions

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

| **Ions that form** *Inoluble* Compounds | **Exceptions** |
|---|---|
| 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 ($NH_4^+$) |

Helpful hints about **Table F.** Group 1 ions appear a lot. Make sure you look at a periodic table to know which elements belong to Group 1.

Don't assume.
Make sure to look at a periodic table to know which elements are **halides** "Group 17"

**Use Table "F" to determine whether the following compounds are soluble or insoluble.**

1. lithium chloride
2. sodium sulfide
3. copper II hydroxide
4. silver chloride
5. zinc chloride
6. copper II nitrate
7. iron III sulfide
8. chromium VI chromate
9. aluminum carbonate
10. ammonium phosphate
11. zinc sulfide
12. lead II acetate

# Colligative Properties

Colligative properties describe the behavior of solutions based on the **number**, rather than the type of solute particles.

What is meant by colligative properties? ________________________________________

**The number of solute particles** affects the following properties of solutions:

- Melting/freezing point
- Boiling/condensation point
- Vapor pressure

In our chemistry class we will focus on what happens to **melting/freezing** and **boiling/condensation** points of a water solution as particles of solute are added to solution. As you recall, the normal melting/freezing point of pure water is 0°C. The normal boiling/condensation point of pure water is 100°C. Notice the range between melting/freezing and boiling/condensation is 100°C.

> Remember! Melting point and freezing point for any substance are the same temperature.

**Adding solute particles to water causes:**

- The melting/freezing point to decrease (become lower).
- The boiling/condensation point to increase (become higher).

If solute particles are added the, result will be to expand the range between the melting and boiling points. This means that the boiling point will be higher and the melting point will be lower. The more particles added, the greater the range. Simple, right? ☺

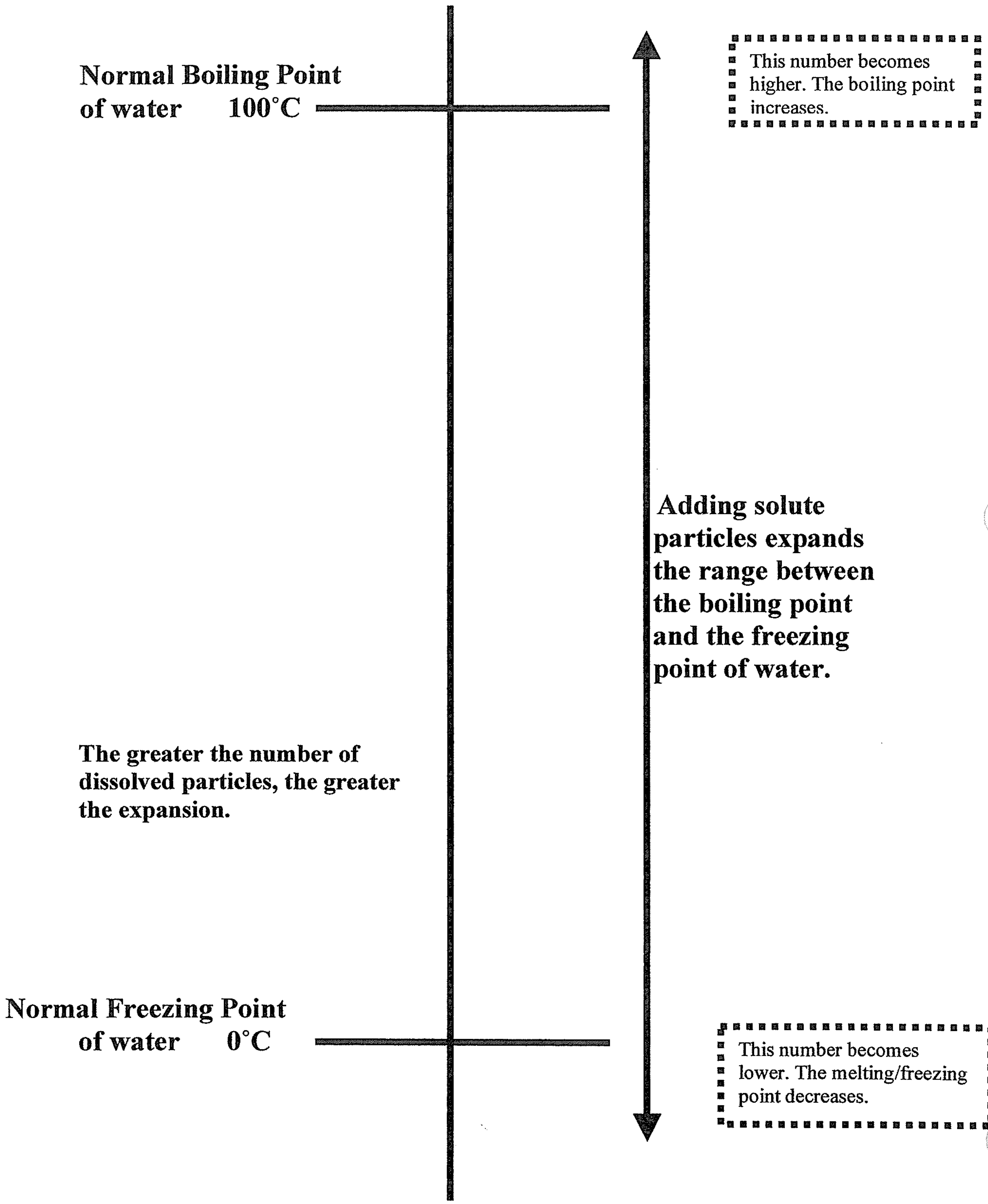
Consider this. If one mole of glucose (molecular) is dissolved into water, it will result in one mole of solute particles. However; if one mole of sodium chloride "NaCl" (ionic) is dissolved in water, it will result in 2 moles of solute particles. Why? Because what happens to soluble ionic compounds when they are dissolved in water? They become hydrated ions (ions surrounded by water molecules). One mole of sodium chloride becomes one mole of sodium ions and one mole of chloride ions, for a total of 2 moles of particles. Dissolving a mole of NaCl will expand the range more than one mole of glucose. The more dissolved particles, the greater the range between the melting and boiling point.

Ionic compounds that provide many ions will expand the temperature range even greater.

**Ex: A mole of $Ba(OH)_2$ provides one mole of $Ba^{+2}$ ions and 2 moles of $(OH)^-$ ions for a total of 3 moles of particles.**

**Ex: A mole of $AlCl_3$ provides one mole of $Al^{+3}$ ions and 3 moles of $Cl^-$ ions for a total of 4 moles of particles.**

**Solute particles (dissolved particles)** expand the range between the freezing and boiling point of water. This means that the boiling point increases (goes up) and the freezing point decreases (goes down). The nature (type) of particles is not important, only the number of particles.



**Question:** State how many moles of solute particles will be provided by the dissolving of 1 mole of the following compounds and state which dissolved compound will expand water's range between freezing and boiling the greatest.
Of course, if you don't know how to identify ionic compounds or correctly write an ionic formula you won't be able to get the job done.

**Sodium chloride**

**Glucose**

**Aluminum sulfate**

Check Table F to see if an ionic compound is soluble. If it is not soluble there will be no ions. Don't assume.

**Carbon dioxide**

**Ammonium hydroxide**

**Calcium chloride**

**Magnesium hydroxide**

It does not matter what the particles are. They can be molecular, ionic, solid, liquid, or gas. The # of dissolved particles determines the change in melting and boiling point, not what they are made of.

**Questions:** A mole of which substance will raise the boiling point of 3 liters of water the most? Why?

**Calcium chloride** **Carbon dioxide** **Aluminum chloride**

# Electrolyte Solutions

**Electrolyte:** a material that dissolves in water to give a solution that conducts electricity.

Describe: electrolyte ______________________________________________

______________________________________________________________

To be an electrolyte a substance must;
**a) dissolve (be soluble) Check Table F for solubility of ionic compounds**
**b) produce ions in solution**

To be an electrolyte a substance must: a) ________________________________

and b) __________________________________________________________

The greater the concentration (molarity) of ions the greater the ability to conduct electricity.

What is the relationship between ion concentration and electrical conductivity of a solution?

______________________________________________________________

**Use the same skills you learned to determine the # of ions produced by dissolving that you used for colligative properties to compare conductivity of solutions.**

Ex: A 2.0 M KI solution will conduct electricity better than 0.2 M KI solution because it has a greater concentration of ions.

**Questions:** Which solution will be a better conductor of electrical current?

a) 2.0 M solution of NaCl or a 1.5 M solution of NaCl? **Why?**

b) 1.0 M solution of $CaCl_2$ or a 1.0 M solution of $AlCl_3$? **Why?**

c) 1.0 M solution of glucose or a 1.0 M solution of NaCl? **Why?**

d) 1.0 M solution of sodium nitrate or 1.0 M solution of magnesium sulfide? **Why?**