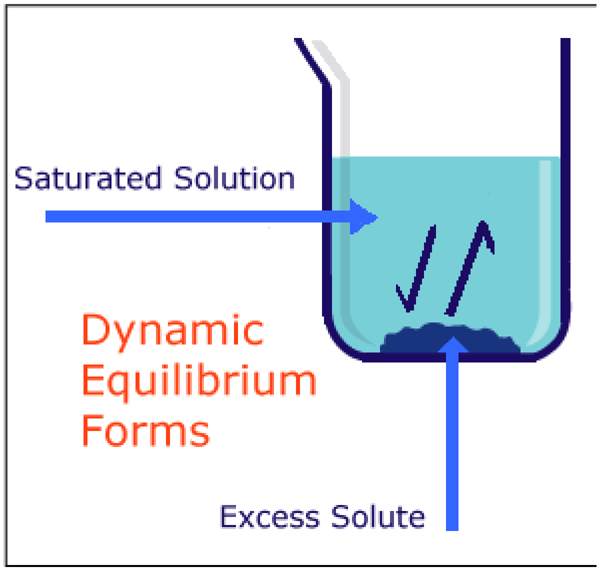
Solutions BASICS

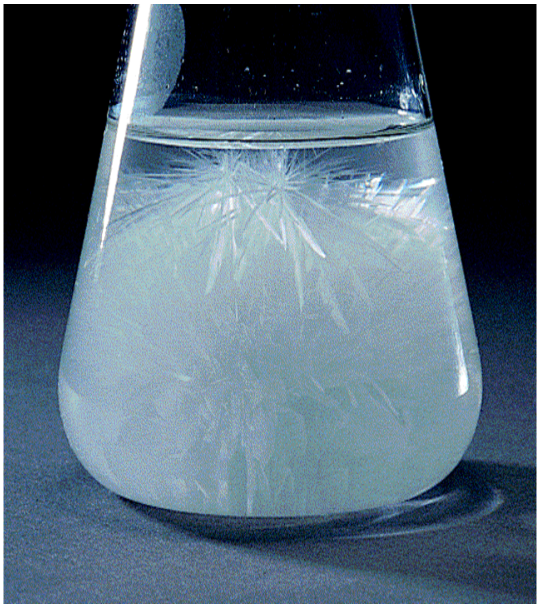
In this section of chemistry we'll be examining solutions, how they form, how to measure their strength, their properties, and how to dilute them exactly to get new solutions of lesser concentration and volume. Then we'll study about changing Colligative properties of water with dissolving particles into it. We'll examine the concept of parts per million for very un-concentrated solutions, finally, we'll do the math with this.

Solutions are homogeneous mixtures containing a solute in a solvent. We most often think of them as wet, with water as the solvent. Other liquids can be solutes as well. Gases can mix homogeneously which makes a gaseous solution, and we could even melt metals or other solids and stir them together. When they cool, technically speaking they are solid solutions (like steel). For now we'll stick to the "wet" solutions.

Solutions can be **saturated**, holding as much solute in a given volume of solvent as possible. At some point there is just no more room in the solvent and added solute cannot be held, so it falls to the bottom of the container. Although a saturated solution is "maxed out", excess solute continues to dissolve into solution while solute falls out of solution –
a dynamic equilibrium is formed. The rate of dissolving is equal to the rate of precipitation. It's a "full" solution, but it's not stuck, rather it's constantly changing while the amount of solute is constant.

An **unsaturated solution** has room to hold more solute. You can add as much solute as you want, and the solution will allow it to dissolve until it reaches the saturation level.



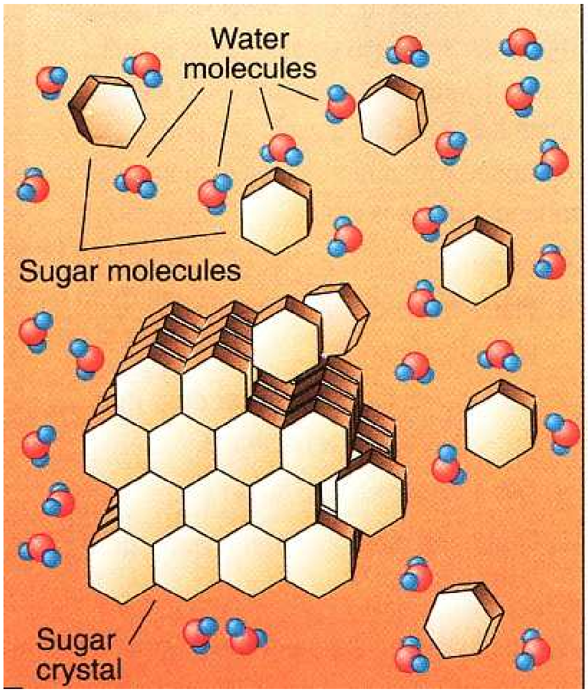


A **supersaturated solution** is one that is more highly concentrated than is normally possible under given conditions of temperature + pressure. Usually you heat up the solvent, saturate it with solute, then cool it to a lower temperature which would not normally be able to contain that amount of solute. If you add some "seed" crystals of solute to this super saturated solution, the excess will collapse out onto these seeds, forming larger crystals. This photo shows the crystallization of excess solute after the seeding.

When these bonds form, energy is released.

Formation of Solutions…

When a crystal of sugar (or other polar compound) is put into the polar solvent water, the crystal is "attacked" by the water molecules. The water molecules surround the sugar molecules, carrying them off molecules of the crystal into solution.



Of course, molecules are too small to see, so the visible crystal is soon invisible to the eye as it's broken into billions of molecules too small to see. At some point the solvent cannot hold a single molecule more, so as more sugar dissolves, some other sugar molecules will precipitate out of solution at the same rate.

<u>Like dissolves like</u> is our solution mantra; polar solvents such as water can only dissolve polar molecular compounds, or most ionic compounds.

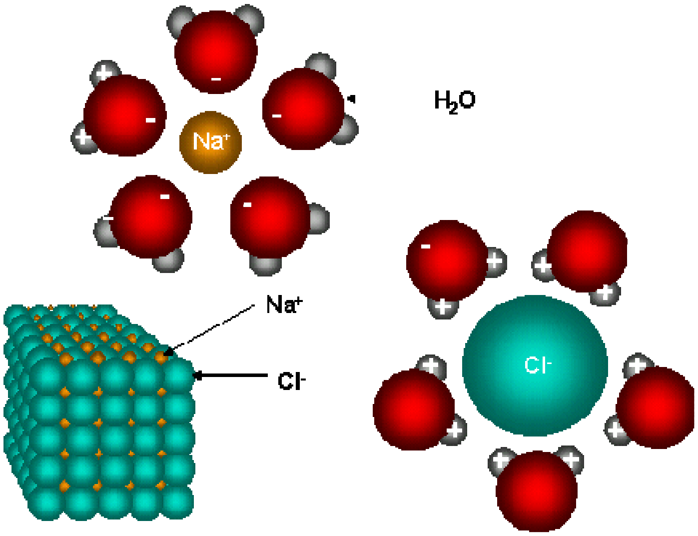
Non-polar compounds can not mix with polar solvents.

At right is oil sitting atop water. The polar water cannot mix with the nonpolar oil. The oil floats because it's less dense. It doesn't mix because: <u>Like Dissolves Like</u> is always true.

When ionic compounds are put into a polar solvent like water, they (usually) are dissociated or ionized into ions. The water molecules surround them as shown below. Solubility exceptions exist on table F!

In the picture below, note how the + side of the water molecules (hydrogen) surround the negative chloride anions. The oxygen, with their — charges, surround the positive sodium cations.

The solvent will dissolve solute until saturated, then the dynamic equilibrium will form.



Remember what an electrolyte is? It's a solution that can conduct electricity. Solutions with ions dissolved can conduct electricity, but solutions with dissolved molecules like sugar cannot conduct. The more ions, the better the conduction. The less ions, the weaker the conduction.

Acids are special chemical compounds in aqueous solutions that appear to be molecular compounds like sugar (no metals), which they are, but they do form ions (we'll learn about acids and bases soon enough).

The CONCENTRATION of solutions.

One of the coolest concepts in chemistry is MOLARITY, the measure of how concentrated a solution is. Molarity can best be described as the molar concentration of a solution, expressed as the number of moles of solute per liter of solution. The formula is:

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

The formula is set up as moles divided by LITERS of solution but any volume of a solution can be made, and its CONCENTRATION will be measured by this formula.

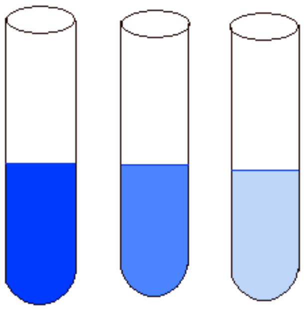
For example…

A 1.0 Molar aqueous solution of HCl could be made by putting 1.0 moles HCl into 1.0 Liters of H_2O.

Or, the same strength or concentration solution could be made with 0.25 moles HCl and 250 mL water.

In fact, an infinite number of combinations of moles to volume exist to make the same concentration of HCl solution.

These three tubes represent 3 different solutions of the SAME compound, but at different concentrations. The darkest one, on the left, would have the HIGHEST MOLARITY or greatest concentration. The one on the far right the LOWEST MOLARITY or least concentration.



A problem may look like this:

What is the concentration of an aqueous solution of KCl containing 370 grams KCl dissolved into 2.5 liters water?

Using the formula above for molarity, we figure this way…

$\text{Molarity} = \frac{\#\text{moles KCl}}{\text{liters of solution}}$	$370\text{ g KCl} \times \frac{1\text{ mole KCl}}{74\text{ grams KCl}} = 5.0\text{ moles KCl}$	$M = \frac{5.0\text{ moles KCl}}{2.5\text{ Liters}}$ $M = 2.0$ molar solution

Making a solution from Scratch.

How do you prepare a 1.00 M of $NaCl_{(AQ)}$ solution of 3.00 Liters in volume?

Start with the molarity formula, putting in the data you have, solving for moles of solute (here that's the NaCl).

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

$$\frac{1.00\ \text{M}}{1} = \frac{\text{\# moles NaCl}}{3.00\ \text{Liters}} = 3.00\ \text{moles NaCl}$$

$$\frac{3.00\ \text{moles NaCl}}{1} \times \frac{58\ \text{grams NaCl}}{1\ \text{moles NaCl}} = 174\ \text{g NaCl}$$

So to make this solution, put 174 grams of NaCl into a large beaker, then fill it up to 3.00 Liters of total volume with water.

NOTE: do not think for one moment that you can put 174 grams of salt into 3.00 Liters of water! That salt has a small but real volume, and this solution is NOT CORRECT. Do not ever make such a silly mistake! Finish your work with the water filling up to the line! Always solute in, THEN water up to the total volume.

How do you make a 1.75 M $CuCl_{2(AQ)}$ of 245 mL? (start with molarity formula)

$$\frac{1.75\ \text{M}}{1} = \frac{\text{\# moles } CuCl_2}{0.245\ \text{Liters}} = 0.429\ \text{moles } CuCl_2$$

$$\frac{0.429\ \text{moles } CuCl_2}{1} \times \frac{134\ \text{grams } CuCl_2}{1\ \text{moles } CuCl_2} = 57.5\ \text{grams } CuCl_2$$

So, to make this solution, put 57.5 grams of copper (II) chloride into a beaker, then fill with water up to the 245 mL mark. DO NOT PUT 57.5 g $CuCl_2$ into 245 ml Water!

The Molar Dilution Formula

Another formula that we can use is called the dilution formula. We can start out with a concentrated stock solution of known volume and molarity, and use it to make a new solution with a new volume and concentration.

How much of the strong solution is needed to create a new solution as stated?

To do a problem like this we substitute in what we know, and calculate our answer. So... For example, assume you have a lot of a concentrated $CuSO_{4(AQ)}$, of 2.0 Molar strength. How would you dilute this to create a 500. mL $CuSO_4$ solution of only 1.0 Molarity? How much of the strong solution is needed? We'll look at the formula, then we'll do the math.

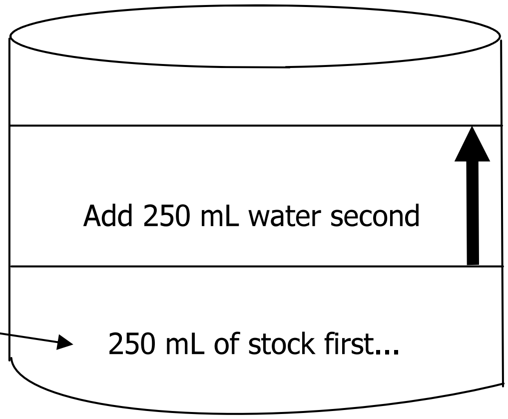
The Molar Dilution formula is…

$$M_1V_1 = M_2V_2$$

$$M_1V_1 = M_2V_2$$

$$(2.0\ M)\ (V_1) = (1.0\ M)(500.\ mL)$$

$$V_1 = 250\ mL \text{ stock solution}$$



This means you will need to add 250 mL of the stronger, original stock solution into a flask, and add enough water to dilute it and fill up to 500. mL solution.

The original solution, the concentrated one, is called a STOCK SOLUTION.

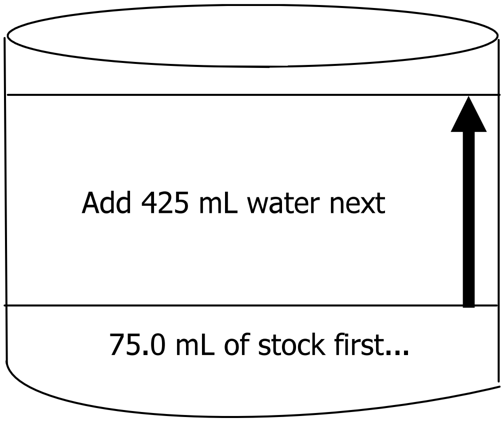
Example 2: Now we'll do a second dilution, to make an even weaker solution of 0.30 M and 500. mL total volume, from the original stock solution.

$$M_1V_1 = M_2V_2$$

$$(2.0\ M)\ (V_1) = (0.30\ M)(500.\ mL)$$

$$V_1 = \frac{(0.30\ M)(500.\ mL)}{2.0\ M}$$

$$V_1 = 75.0\ mL$$

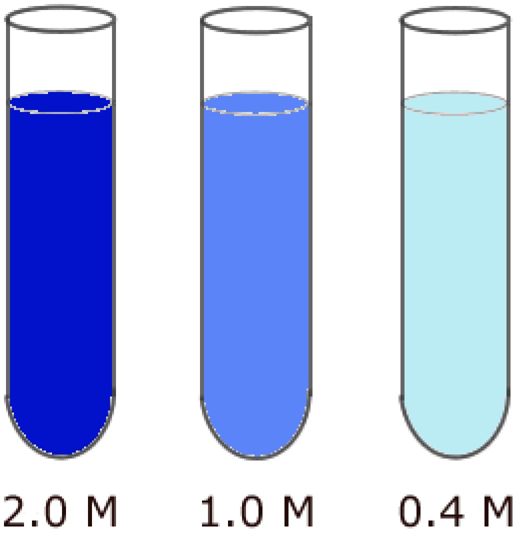


75.0 mL of 2.0M $CuSO_{4(AQ)}$ + 425 mL H_2O = 500 mL of 0.30M $CuSO_{4(AQ)}$

These three tubes represent the $CuSO_4$ solutions we just made.
First is the 2.0 M stock solution
Second is the 1.0 M diluted solution
Third is the final weakest 0.4 M $CuSO_{4(AQ)}$.

All made of the same $CuSO_4$, but of different concentrations or strengths.

Aqueous solutions of $CuSO_4$ would be shades of blue, the darkness of solution would depend upon the concentration or molarity of the solutions.



Another example:

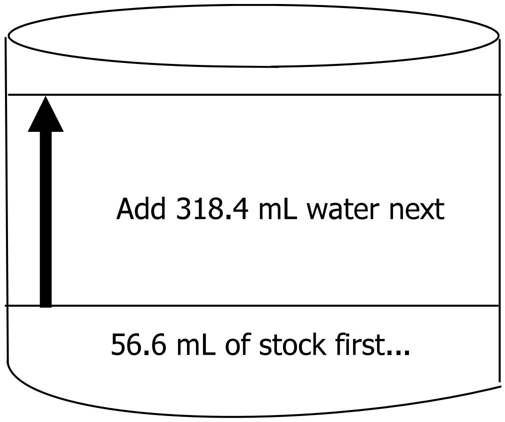
You have a 5.00 M stock NaCl solution. You want to prepare a 375. mL salt water solution of 0.755 M concentration. When you start with a stock solution, you need to dilute it, with the dilution formula:

$$M_1V_1 = M_2V_2$$

$$(5.00\text{ M})(V_1) = (0.755\text{ M})(375\text{ mL})$$

$$V_1 = \frac{(0.755\text{ M})(375\text{ mL})}{5.00\text{ M}}$$

$$V_1 = 56.6\text{ mL of stock solution}$$



That means, you need 56.6 mL of the concentrated stock salt water solution and you need to dilute it then with enough water to reach the 375 mL mark, which is about 318.4 mL (disregarding SF).

Note: to make a solution from scratch you use the molarity formula. To make a solution from an existing stock solution, use the dilution formula.

No matter, always draw a picture of an empty beaker, and “show making the solution” so you can see clearly what you are doing.

Colligative Properties of Solutions

These are physical properties that can change depending upon how much solute is dissolved into a liter of the solution. They include boiling point, freezing point, & vapor pressure. These three different properties get adjusted by the solute mixed into the water. If you dissolve particles (ions or polar molecules) into water, you change all of these properties. The more particles in solution, the greater the properties change.

First we need to examine what happens when substances dissolve into water. Molecular compounds, like sugar, dissolve into water, they do not form ions. They are not ionic. When soluble ionic compounds dissolve, the compound ionizes, or it dissociates into ions this way:

$C_6H_{12}O_6$	$C_6H_{12}O_{6(S)} \rightarrow C_6H_{12}O_{6(AQ)}$	1 mole $C_6H_{12}O_6 \rightarrow$ 1 mole of molecules
NaCl	$NaCl_{(S)} \rightarrow Na^{+1}_{(AQ)} + Cl^{-1}_{(AQ)}$	1 mole NaCl → 2 moles of ions
$CaCl_2$	$CaCl_2 \rightarrow Ca^{+2}_{(AQ)} + Cl^{-1}_{(AQ)} + Cl^{-1}_{(AQ)}$	1 mole $CaCl_2 \rightarrow$ 3 moles of ions
$AlCl_3$	$AlCl_3 \rightarrow Al^{+3}_{(AQ)} + Cl^{-1}_{(AQ)} + Cl^{-1}_{(AQ)} + Cl^{-1}_{(AQ)}$	1 mole $AlCl_3 \rightarrow$ 4 moles of ions
AgCl	AgCl → no moles of ions, it's insoluble!	1 mole AgCl = zero particles.

So when things dissolve into water, depending upon what the substance is, the change of colligative properties is not always the same. 1 mole of substance does not always equal one mole of particles.

BOILING POINT ELEVATION

The water boils when it can overcome both the air pressure pressing down on the surface, and the internal hydrogen bonding holding the molecules together. At normal pressure the boiling point of pure water is 373 Kelvin. When polar molecules or ions are dissolved into the water, the water molecules are ALSO attracted to the particles. This creates MORE INTERNAL ATTRACTION, which means it will take more energy to make the water boil (blow apart from itself into the gas phase).

The actual boiling point elevation for water is 0.50 K per mole of particles per liter of solution. For every mole of particles, the boiling point goes up by 0.50 Kelvin.

Example 1. What is the boiling point of a 1.0 Molar $NaCl$ solution of one liter?

Start boiling point	Boiling point elevation	New boiling point
373 K	+ (2 x 0.50 K) =	373 + 1 = 374 K

Example 2. What is the boiling point of a 1.0 Molar $CaCl_2$ solution of one liter?

Start boiling point	Boiling point elevation	New boiling point
373 K	+ (3 x 0.50 K) =	373 + 1.5 = 374.5 K

Example 3. What is the boiling point of a 2.0 Molar $CaCl_2$ solution of one liter?

Start boiling point	Boiling point elevation	New boiling point
373 K	+ (6 x 0.50 K) =	373 + 3 = 376 K

Did you see that one? Each mole of calcium chloride forms 3 moles of ions.
Here the solution contains 2 moles of calcium chloride, each forms 3 moles of ions, so 2 X 3 = 6

Example 4. What is the boiling point of a 4.0 Molar $AlCl_3$ solution of one liter?

Start boiling point	Boiling point elevation	New boiling point
373 K	+ (16 x 0.50 K) =	373 + 8 = 381 K

Did you see that one? Each mole of aluminum chloride forms 4 moles of ions.
Here the solution contains 4 moles of calcium chloride, each forms 4 moles of ions, so 4 X 4 = 16

Example 5. What is the boiling point of a 2.5 Molar $CaBr_2$ solution of one liter?

Start boiling point	Boiling point elevation	New boiling point
373 K	+ (7.5 x 0.50 K) =	373 + 3.75 = 386.75 K

Did you see that one? Each mole of calcium bromide forms 3 moles of ions. Here the solution contains 2.5 moles of calcium chloride, each forms 3 moles of ions, so 2.5 X 3 = 7.5

FREEZING POINT DEPRESSION

The freezing point is also very affected by dissolved particles. The difference is that the freezing point requires colder temperatures to freeze around the annoying, cluttered ions or molecules that are "in the way" of the hydrogen bonding. Water only freezes when the hydrogen bonding is stronger than the kinetic energy that the particles have (the temperature more or less is the KE).

The freezing point depression for water is that each mole of particles depresses the freezing point by 1.86 Kelvin (that's a lot!).

$C_6H_{12}O_{6(S)} \rightarrow C_6H_{12}O_{6(AQ)}$	1 mole $C_6H_{12}O_6 \rightarrow$ 1 mole of molecules
$NaCl_{(S)} \rightarrow Na^{+1}_{(AQ)} + Cl^{-1}_{(AQ)}$	1 mole NaCl → 2 moles of ions
$CaCl_2 \rightarrow Ca^{+2}_{(AQ)} + Cl^{-1}_{(AQ)} + Cl^{-1}_{(AQ)}$	1 mole $CaCl_2 \rightarrow$ 3 moles of ions
$AlCl_3 \rightarrow Al^{+3}_{(AQ)} + Cl^{-1}_{(AQ)} + Cl^{-1}_{(AQ)} + Cl^{-1}_{(AQ)}$	1 mole $AlCl_3 \rightarrow$ 3 moles of ions
AgCl → no moles of ions, it's insoluble!	1 mole AgCl = zero particles.

Example 6. What is the freezing point of a 1.0 Molar NaCl solution of one liter?

Start freezing point	Freezing point depression	New boiling point
273 K	– (2 x 1.86 K) =	273 – 3.72 = 269.28 K

Example 7. What is the freezing point of a 2.0 Molar $CaCl_2$ solution of one liter?

Start freezing point	Freezing point depression	New boiling point
273 K	– (6 x 1.86 K) =	273 – 11.16 = 261.84 K

Did you see that? Each mole of calcium chloride forms 3 moles of ions. Here there are 2 moles, so, 2 x 3 = 6 moles of particles in solution.

Example 8. What is the freezing point of a 4.0 Molar $AlCl_3$ solution of one liter?		
Start freezing point	Freezing point depression	New boiling point
273 K	– (16 x 1.86 K) =	273 – 29.76 = 243.24 K
Did you see that? Each mole of aluminum chloride forms 4 moles of ions. Here there are 4 moles, so, 4 x 4 = 16 moles of particles in solution.		

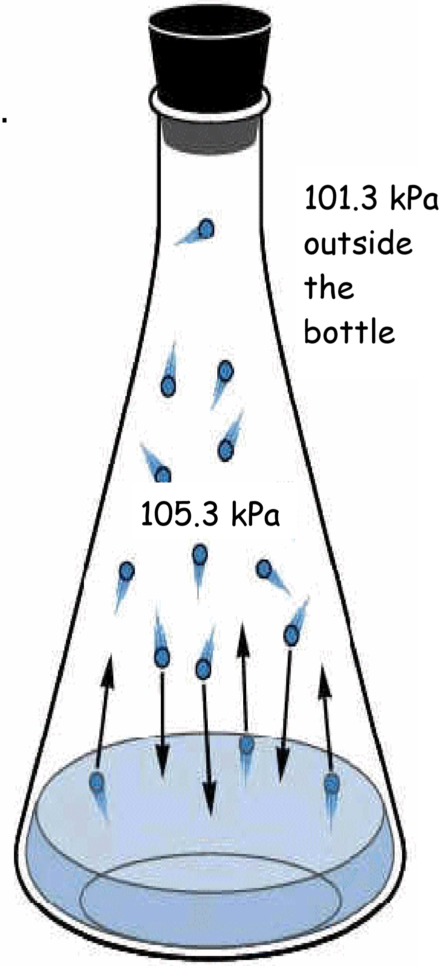
Just like with the boiling point elevation, it's so important to realize how many moles of particles you get when you dissolve molecular compounds, or salts into water. Molecular compounds DO NOT IONIZE or dissociate. One mole of sugar gives you 1 mole of molecules. One mole of ethanol alcohol gives one mole of particles.
These molecular compounds ARE NOT IONIC so they can't possible ionize.

Ionic compounds do ionize, each into a particular number of ions. Each mole of ions depresses the freezing point by 1.86 Kelvin.

VAPOR PRESSURE ADJUSTMENT

The vapor pressure is shown in table H on the reference tables.
The vapor pressure is THE EXTRA PRESSURE ADDED
TO A CLOSED SYSTEM BY THE EVAPORATION OF A LIQUID.

Room temperature water (25°C) has a vapor pressure of about 4 kPa. That means inside that bottle at right, if the starting air pressure inside the bottle was 101.3 kPa (normal), the water evaporating will add to it by about +4 kPa.
That makes the pressure in the bottle about 105.3 kPa.



Water doesn't evaporate well because of all the hydrogen bonding it has. That means water has a low vapor pressure. Adding any solute to water only increases the internal attraction, making it harder to evaporate.

In our class we will only know that any solute in water decreases the vapor pressure
(makes it evaporate less well).
More concentrated aqueous solutions have
lower vapor pressure compared to dilute aqueous solutions. We'll do NO MATH for the vapor pressure in this part of chem.

PARTS PER MILLION

Some solutions are so very dilute that the molarity becomes a super small decimal that our brains can't make easy sense of. For example, when you add just 1.0 moles of NaCl into a swimming pool sized 43,000 Liter solution, the molarity works this way:

$\text{Molarity} = \dfrac{\text{moles of solute}}{\text{Liters of solution}}$	$\text{Molarity} = \dfrac{0.250 \text{ moles NaCl}}{43{,}000 \text{ Liters}}$	M = 0.00000581 M or it's 5.81×10^{-6} Molar NaCl solution which is silly small

This number is accurate, but 581 hundred millionths molar is just outside normal thinking zones.

Another way to do concentration is parts per million.

$$\text{PPM} = \frac{\text{Grams solute}}{\text{Grams water}} \times 1{,}000{,}000 = \frac{58 \text{ g NaCl}}{43{,}000{,}000 \text{ g } H_2O} \times 1{,}000{,}000 = \mathbf{1.35\ PPM\ salt}$$

1.35 PPM is a number you can more easily wrap your head around, although it is perfectly equivalent to the silly small molarity just above it.

Example A

If there is 0.125 grams of mercury dissolved into 101. liters of sea water, concentration is parts per million?

$$\text{PPM} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 1{,}000{,}000$$

$$\text{PPM} = \frac{0.125 \text{ g Hg}}{101{,}000 \text{ g water}} \times 1{,}000{,}000 = 1.24 \text{ parts per million}$$

This means: 1.24 parts Hg per million parts water

Example B From an old regents exam was this problem…

What is the concentration of a solution in parts per million if 0.02 grams Na_3PO_4 is dissolved into 1000 grams water?

A. 20 PPM B. 2 PPM C. 0.2 PPM D. 0.02 PPM

ANSWER

$$\text{PPM} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 1{,}000{,}000$$

$$\text{PPM} = \frac{0.02\text{ g }Na_3PO_4}{1000\text{ g water}} \times 1{,}000{,}000 = \text{20 parts per million (choice A)}$$

The regents will sometimes uses just one significant figure in a problem, they are trying to get across the concept, and not the math. You always pick the BEST possible choice, which sometimes won't earn full credit in our class. SF count, but the regents breaks that rule often.

Bits and Pieces

To make a proper solution, of perfect volume, there is only one way to proceed. First get a special flask with a line that shows exactly a particular volume (often 1.00 L). Put the solute in first. Then fill with pure water up to the line.
This is the ONLY WAY to make a perfect solution. You can't just add solute into the solvent, it will affect the volume (in a small but measurable way).

When an ionic compound (NaCl, KCl, etc.) is dissolved into water it forms an ionic solution. It has free ions floating in the water. This is a homogeneous mixture. The more ions in solution, the better electrical conduction that occurs. Fewer ions means a lesser electrical conduction.

If you melt an ionic compound like $NaCl_{(L)}$ or $CuBr_{(L)}$ it will be super duper hot. It will also be able to conduct electricity because the ions are loose, almost like in an aqueous solution. This is weird, it would be way too hot to handle in most colleges and impossible in high school, but it would conduct. Ionic compounds that do not ionize in water are NOT electrolytes, but they can conduct electricity is their liquid or melted states.

Electrolytes are solutions with ions in them (soluble ionic compounds), and they can conduct electricity. Electrolytes are always able to conduct electricity. The more ions in solution, the better the electricity flows.

Ionic compounds in the solid form CANNOT conduct electricity because there are no loose ions, and no loose electrons (as with metals and metallic bonds). Solid ionic compounds are called electrolytes only if they are soluble in water. Electrolytes can be solutions that conduct, or (strangely enough) solids that *would form* soluble ionic solutions. Insoluble ionic compounds like AgCl, or molecular compounds are never electrolytes, and cannot conduct electricity.