Lesson 01 and 02: Solutions, Solubility and Conductivity

01 What is a Solution?

Before we can talk about solubility it is important to look at the term solution. What is a solution?

A solution is a homogeneous mixture (uniform in composition) composed of two or more substances. In such a mixture, a solute is dissolved in a solvent. A solute is the substance being dissolved and a solvent is the substance being dissolved into.

Common Solutions		Solute: Gas	Solute: Liquid	Solute: Solid
Solvent	Gas	Oxygen and in nitrogen (air).	Water vapour in air.	Naphthalene slowly sublimes in air.
Solvent	Liquid	Carbon dioxide in water.	Ethanol in water; various hydrocarbons in each other.	Sucrose in water; sodium chloride in water; gold in mercury, forming an amalgam.
Solvent	Solid	Hydrogen dissolves rather well in metals; platinum has been studied as a storage medium.	Hexane in paraffin wax, mercury in gold.	Steel, aluminum, other metal alloys.

02 What is Solubility?

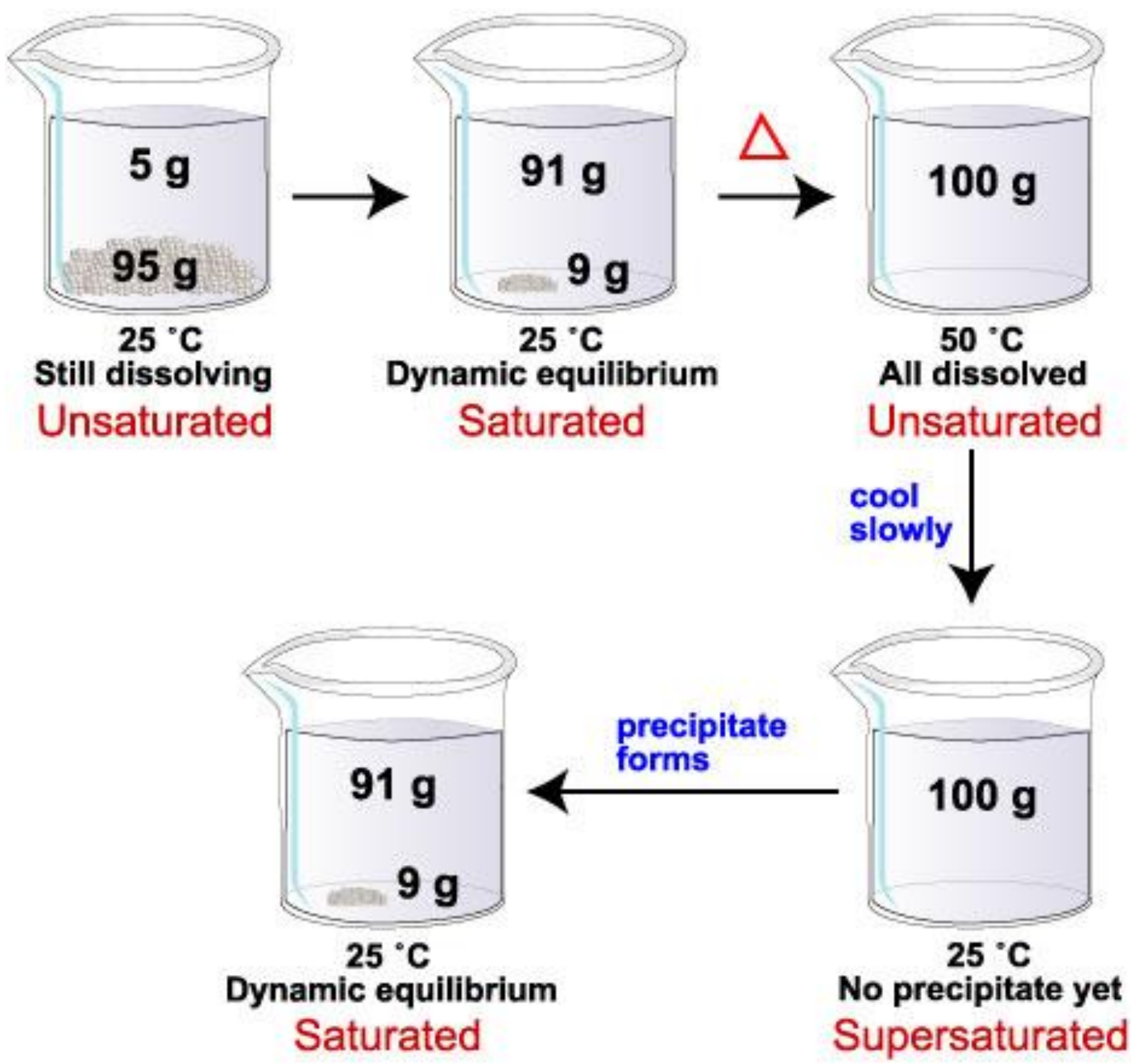
Solubility is generally considered to be the maximum amount of a substance (the solute) that can be dissolved into a given amount of solvent, at a particular temperature.

Once a solution has reached the limit of the solute's and solvent's solubility, the solution is said to be saturated, meaning that it can hold no more solute. If additional solute is added to a saturated solution, the extra solute will settle out, forming a separate layer.

If a solution has not dissolved the maximum amount of a particular substance it is said to be unsaturated.

If a solution has dissolved more than the maximum amount of a particular substance (than it would normally) it is said to be supersaturated.

Saturated, Unsaturated and Supersaturated Solutions



03 Distinguishing Aqueous Solutions: Conductivity

There are two main types of aqueous solutions. Each solution contains compounds that are held together differently. Each solution can be distinguished from the other based on conductivity…

Two Main Types of Solutions	
Solutions Containing Covalent Compounds	Solutions Containing Ionic Compounds
covalent bonding	ionic bonding
sharing of electrons	transfer of electrons
gases, organic compounds, diatomic molecules…	acids, bases and salts
CO_2, NO_2, C_2H_2, CH_4, $C_6H_{12}O_6$, N_2, O_2, H_2, S_8…	HCl, NaOH, NaCl…

Covalent and Ionic Compounds

Covalent compounds are compounds composed of a non-metal and a non-metal. Covalent compounds form molecular solutions…

$$C_2H_2(g) \rightarrow C_2H_2(aq)$$

$$CO_2(g) \leftrightarrow CO_2(aq)$$

Ionic compounds are typically composed of a metal (positive charge) and a non-metal (negative charge), but can be non-metal (positive charge) and non-metal (negative charge). Ionic compounds form ionic solutions…

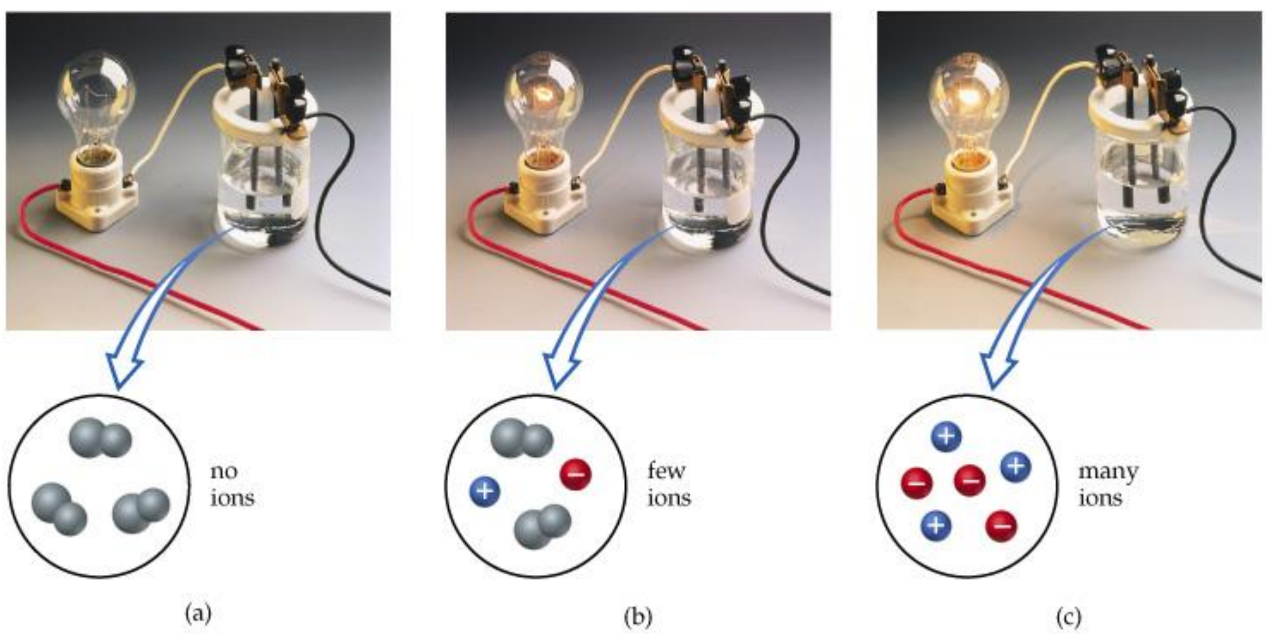
$$\text{Acids: } HCl(l) \rightarrow H^{+}(aq) + Cl^{-}(aq)$$

$$\text{Bases: } NaOH(s) \rightarrow Na^{+}(aq) + OH^{-}(aq)$$

$$\text{Salts: } NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$$

Conductivity of Covalent and Ionic Solutions

How can one determine whether or not a solution contains an ionic substance or a covalent substance? If a solution contains ions the solution will conduct electricity. If not, it will not.



Concentration and Conductivity

Keep in mind that concentration plays a part in conductivity as well.

If a solution is ionic and the concentration is low the conductivity will also be low. If the solution is ionic and concentration is high the conductivity will be high.

If a solution is molecular, the conductivity will always be low, no matter the concentration.

Example	
Which of the following solutions are conductive and which are not?	
$BaBr_2$(aq)	
CO_2(aq)	
KBr(s)	
K_2CrO_4(aq)	
$NaNO_3$(s)	
$C_6H_{12}O_6$(aq)	

Example	
Which of the following pairs of solutions would you expect to be the most conductive?	
1 M $BaBr_2$(aq)	0.5 M $BaBr_2$(aq)
0.2 M K_2CrO_4(aq)	2 M K_2CrO_4(aq)
1 M K_2CrO_4(aq)	1 M $FeCl_3$(l)
1 M H_3PO_4(aq)	1 M HNO_3(aq)

Example
Describe what you would observe if you were to test the following two solutions for conductivity: (1) a solution containing a small number of ions, and (2) a solution containing a large number of ions.

Lesson 03, 04 and 05: Intermolecular Forces and Properties of Substances

01 Background

When a substance dissolves, boils, or melts the molecules involved must be…

- driven apart, and
- stay apart

In order to stay apart, the forces of attraction between molecules must be overcome.

Two question must then be answered…

- what forces are present?
- how strong are they?

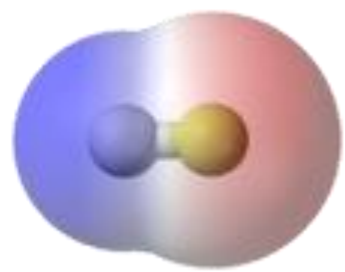
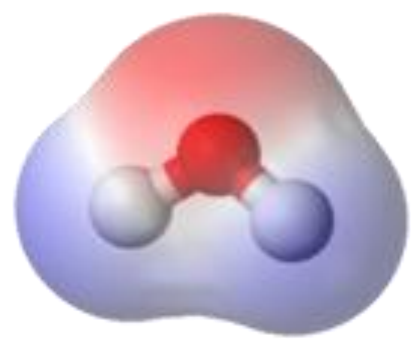
02 Intermolecular Forces Review

There are two types of intermolecular forces of concern to us...

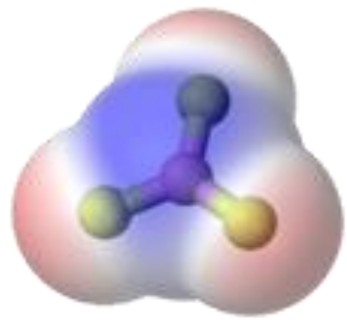
- permanent dipole - permanent dipole force, such as the hydrogen bond
- instantaneous dipole - induced dipole force such as the London dispersion force

Of the two types of intermolecular forces…

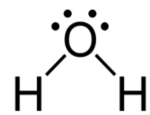
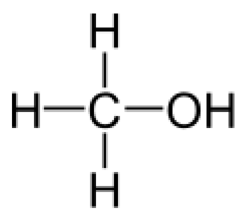
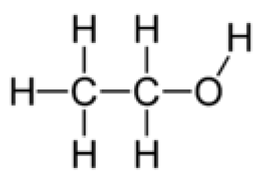
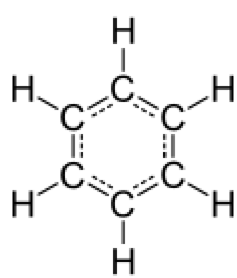
- permanent dipole - permanent dipole forces involve molecules that are polar

Polar Molecules	
Sharing of Electrons	unequal sharing of electrons between bonded atoms
Arrangement of Atoms	asymmetrical arrangement of bonded atoms
Examples (HF and H_2O)	 

- instantaneous dipole - induced dipole force involve molecules that are not polar but become polar

Non-Polar Molecules	
Sharing of Electrons	equal or unequal sharing of electrons between bonded atoms
Arrangement of Atoms	symmetrical arrangement of bonded atoms
Example	

Example 01

Classify Compounds as Polar or Nonpolar		
Compound	Structure	Polar?
water		YES
methanol		YES
ethanol		YES
benzene		NO

03 Intermolecular Forces and Properties of Substance

Properties of substances such as melting, boiling or dissolving require that intermolecular forces be overcome. This can be easy or difficult depending on the strength and combination of intermolecular forces involved...

<table>
<tr><th colspan="3">Relative Strengths of Intermolecular Forces (kJ/mol)</th></tr>
<tr><td>Permanent Dipole – Permanent Dipole</td><td colspan="2">Instantaneous Dipole – Induced Dipole</td></tr>
<tr><td>Hydrogen Bond</td><td>All Other PD - PD</td><td>London Dispersion Forces</td></tr>
<tr><td>10-40</td><td>5-25</td><td>0.05-40</td></tr>
</table>

<table>
<tr><th colspan="5">Example 02</th></tr>
<tr><th colspan="5">Boiling Points and Intermolecular Forces</th></tr>
<tr><th>Example</th><th>Boiling</th><th>Number of</th><th>Intermolecular</th><th>Boiling</th></tr>
</table>

	Temp.	Electrons	Force(s) Present	Point
$N_2(l)$	-196C	14	small London dispersion force	high → low
$O_2(l)$	-183C	16	small London dispersion force	
NO(l)	-152C	15	small London dispersion force and dipole force	
$Br_2(l)$	59C	70	large London dispersion force	
ICl(l)	97C	70	large London dispersion force and dipole force	

Example 03				
Melting Points and Intermolecular Forces				
Example	Melting Temp.	Number of Electrons	Intermolecular Force(s) Present	Melting Point

NH_3	-78C	decreasing →	London dispersion forces and hydrogen bonding	high
PH_3	-133C		decreasing London forces →	high → low
AsH_3	-116C			
SbH_3	-88C			

Example 04		
Determine the Highest Boiling Point		
Compound 01	Compound 02	Reasons

H_2O	H_2S	Compound 01, because it contains hydrogen bonding and London forces.
CH_3CH_2OH	CH_3CH_2SH	Compound 01, because it contains hydrogen bonding and London forces.
CH_3CH_3	CH_3NH_2	Compound 02, because it contains hydrogen bonding and London forces.

04 Intermolecular Forces and the Dissolving Process

The dissolving process involves a consideration of the types of intermolecular attractive forces between solute-solute molecules and solvent-solvent molecules…

- A solute WILL DISSOLVE in a solvent if the solute-solvent forces of attraction are great enough to overcome the solute-solute and solvent-solvent forces of attraction.

- A solute WILL NOT DISSOLVE if the solute-solvent forces of attraction are weaker than individual solute and solvent intermolecular attractions

The reason for this is because these intermolecular attractions must be broken before new solute-solvent attractive forces can be established.

Therefore, in general…

- ions are attracted to the molecules of polar solvents
- non-polar molecules are attracted to non-polar molecules of non-polar solvents

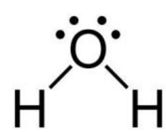
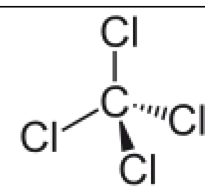
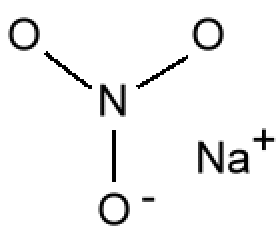
A polar substance is more likely to dissolve in a polar solvent, and a non-polar substance is more likely to dissolve in a non-polar solvent…

- like dissolves like

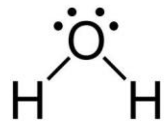
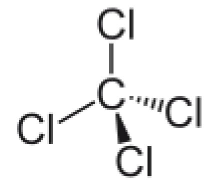
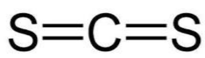
Example 05	
Solubility of Polar and Nonpolar Solutes in Polar Solvents	
	Polar Solvent
Polar Solute	YES
	Permanent dipole – permanent dipole forces and London Dispersion forces present.
Nonpolar Solute	NO
	Only London dispersion forces present. Cannot overcome dipole forces of polar solvent.

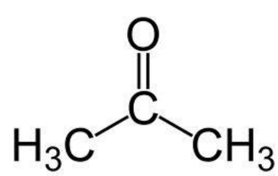
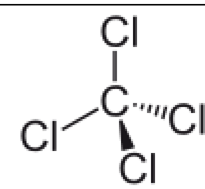
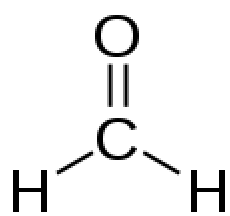
Example 06	
Solubility of Polar and Nonpolar Solutes in Nonpolar Solvents	
	Nonpolar Solvent
Polar Solute	NO

	Only London dispersion forces present. Cannot overcome dipole forces of polar solute.
Nonpolar Solute	YES
	Only London dispersion forces present. No dipole forces to overcome.

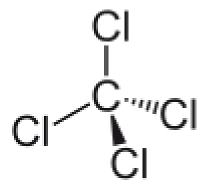
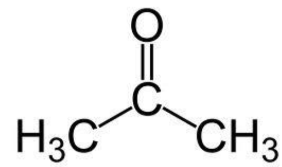
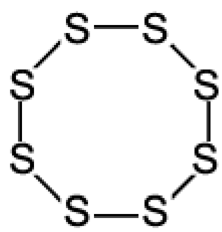
Example 07		
Classify as Soluble or Insoluble		
Solute / Solvent	Water	Carbon Tetrachloride
		
Sodium Nitrate 	Soluble	Insoluble

Example 08		
Classify as Soluble or Insoluble		
Solute / Solvent	Water	Carbon Tetrachloride

		
Carbon disulphide 	Insoluble	Soluble

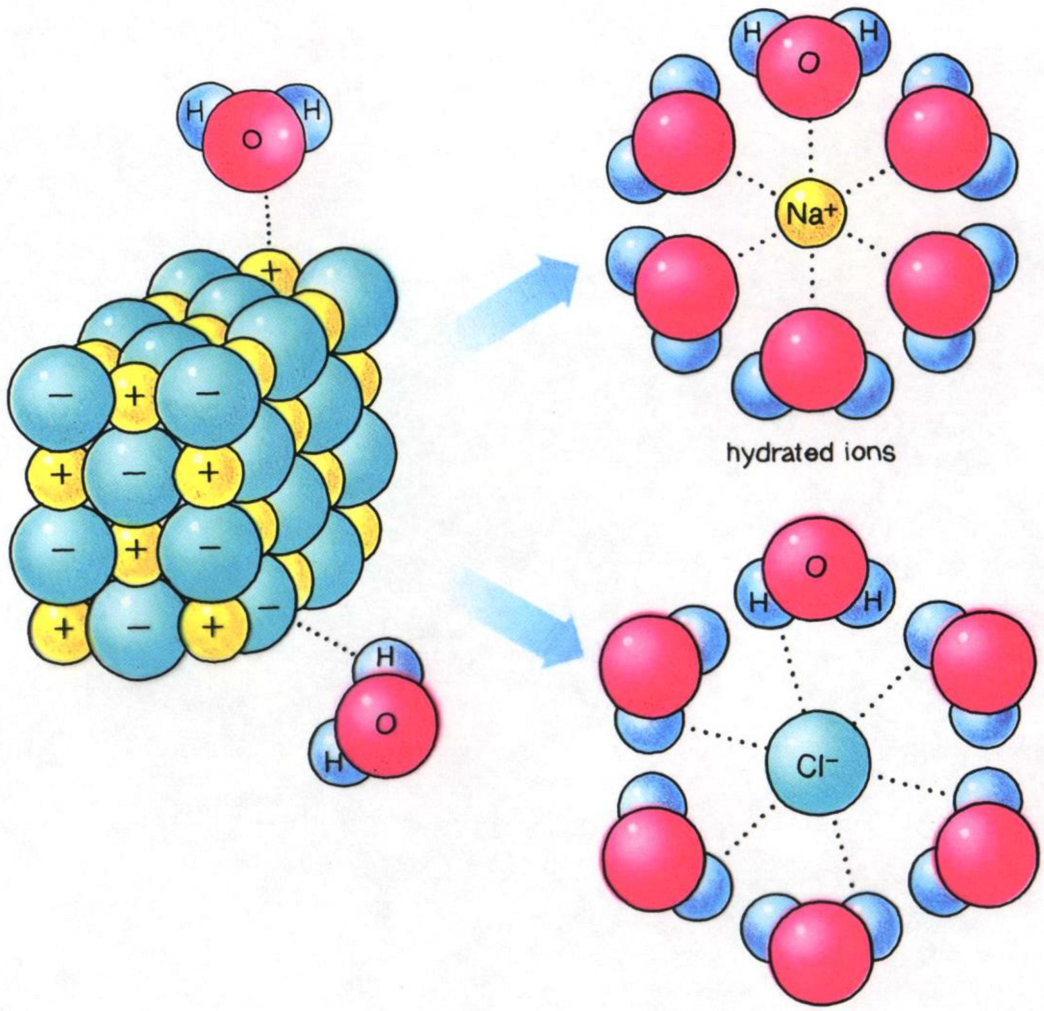
Example 09		
Classify as Soluble or Insoluble		
Solute / Solvent	Acetone 	Carbon Tetrachloride 
Formaldehyde 	Soluble	Insoluble

Example 10		
Classify as Soluble or Insoluble		
Solute / Solvent	Carbon Tetrachloride	Acetone

		
Sulphur 	Soluble	Insoluble

05 Aqueous Solutions: The Dissolving Process

The dissolving process, sometimes called solvation or dissolution, is the process of attraction and association of molecules of a solvent with molecules or ions of a solute. If solute and solvent intermolecular forces are compatible, solute molecules or ions dissolve, spread out, and become surrounded by solvent molecules.



Two terms that describe the formation of ions in solution, and are often confused are...

- dissociation: Refers to the formation of ions from an ionic solid. Remember an ionic solid is composed of a metal and a nonmetal. Ions already exist in an ionic solid. All that is happening is separation of those ions.

$$NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$$

- ionization: Refers to the formation of ions from a covalent or molecular solid. Ions do not exist until the solvent has reacted.

$$CH_3COOH(l) \rightarrow CH_3COO^-(aq) + H^+(aq)$$

Example 11	
Write Dissociation / Ionization Equations	
Compound	Equation
$Ca(OH)_2(s)$	$Ca(OH)_2(s) \rightarrow Ca^{2+}(aq)+2OH^-(aq)$
$HCl(g)$	$HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$
$(NH_4)_2S(s)$	$(NH_4)_2S(s) \rightarrow 2NH_4^+(aq)+S^{2-}(aq)$
$AlCl_3(s)$	$AlCl_3(s) \rightarrow Al^{3+}(aq)+3Cl^-(aq)$

Lesson 06: Concentration of Ions in Aqueous Solutions

01 Ion Concentrations for Solids Added to Water

When determining ion concentrations for solutions it is important to remember the mole ratio concept for balanced chemical reactions...

- determine the balanced chemical reaction for dissociation
- concentrations of ions present depends on the mole ratio between reactants and products

Example 01

Calculate the concentrations of each ion for the following solution… 12.0 grams of $(NH_4)_2CO_3$ in 2.50 liters of water

- first write out the dissociation reaction…

$$(NH_4)_2CO_3(aq) \rightarrow 2NH_4^{+}(aq) + CO_3^{2-}(aq)$$

- second, calculate the concentration of the substance in water before dissociation takes place…

$$\left(\frac{12.0\text{ g }(NH_4)_2CO_3}{2.50\text{ L}}\right)\left(\frac{1\text{ mole }(NH_4)_2CO_3}{96.0\text{ g }(NH_4)_2CO_3}\right) =$$

$$0.0500\text{ M }(NH_4)_2CO_3$$

- lastly, calculate the concentration of all ions

$$\left[NH_4^{\ +}\right] = 2\left(0.0500\text{ M }\ (NH_4)_2\,CO_3\right) = 0.100\text{ M}$$

$$\left[CO_3^{\ 2-}\right] = 1\left(0.0500\text{ M }\ (NH_4)_2\,CO_3\right) = 0.0500\text{ M}$$

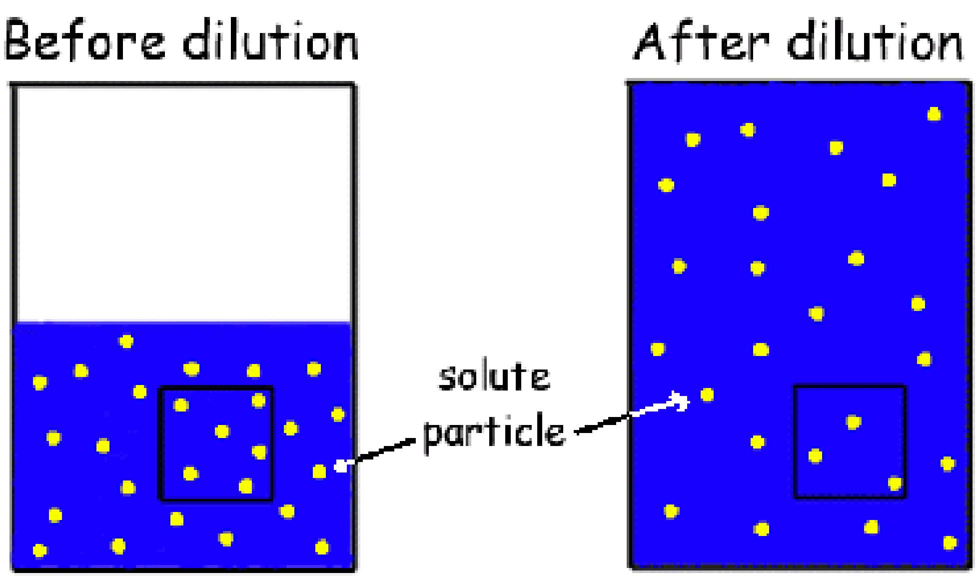
02 Ion Concentrations for Dilutions

For convenience, solutions are either purchased or prepared in concentrated stock solutions which must be diluted prior to use.

When we take a sample of a stock solution we have a certain concentration of molecules in that sample.

Dilution alters…

- the concentration of the solution
- but not the total number of moles in the solution.



One of the standard equations for determining the effects of dilution upon a sample is to set up an equation comparing concentration and volume before and after dilution…

$$M_{before}V_{before} = M_{after}V_{after}$$

$$M_{after} = \frac{M_{before}V_{before}}{V_{after}}$$

Example

A chemist starts with 50.0 mL of a 0.40 M sodium chloride solution and dilutes it to 1000.0 mL (add 950.0ml). What is the concentration of sodium chloride in the new solution?

$$M_{dilute}V_{dilute} = M_{concentrated}V_{concentrated}$$

$$M_{dilute} = \frac{M_{concentrated}V_{concentrated}}{V_{dilute}} = \frac{(0.40M)(50.0ml)}{1000.0ml} = 0.020M$$