<u>Unit 10 - Solution Chemistry</u>

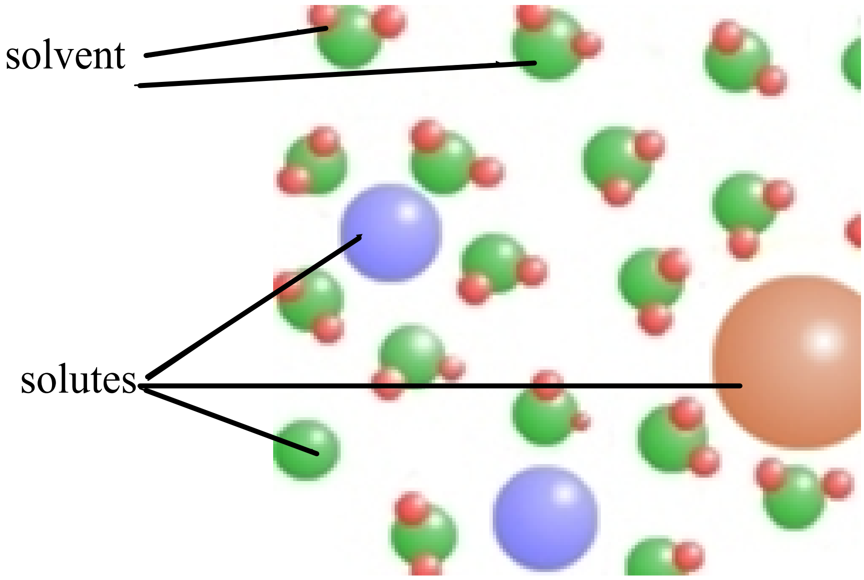
1. Solutions & Molarity

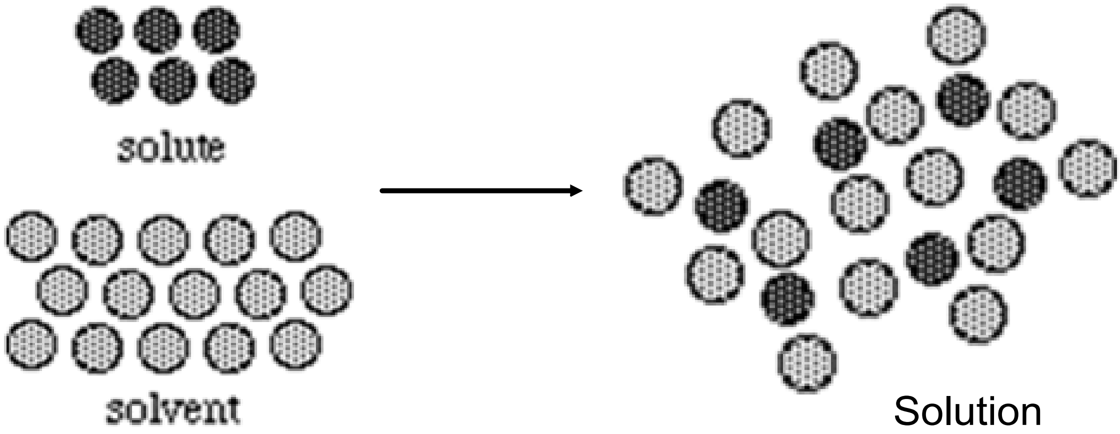
- Solution: a homogeneous mixture of two or more substances
- Said another way -
- a mixture of two or more different types of particles that looks like one substance
- ex: salt water, kool-aid

- Solvent:
- The most abundant component of a solution
- Solute
- The component(s) that are less abundant

In salt water:
Solvent: water **Solute**: salt

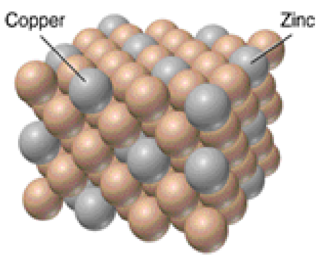
In Kool-Aid:
Solvent: water **Solutes**: sugar, Kool-Aid crystals


solvent
solutes

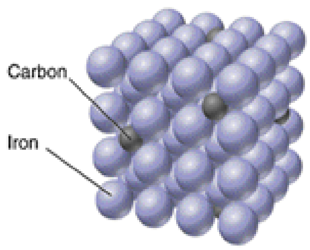

solute
solvent
Solution

- The symbol *(aq)* after a formula – stands for *aqueous* - indicates a solution that has water as the solvent
- ex. $NaCl_{(aq)}$ means “NaCl dissolved in water”

- Solutions can be:
- liquid – solid
- salt water, Kool-Aid
- gas – gas
- air
- liquid – liquid
- water and vinegar
- liquid – gas
- pop
- solid – solid
- metal alloys: bronze, steel, brass



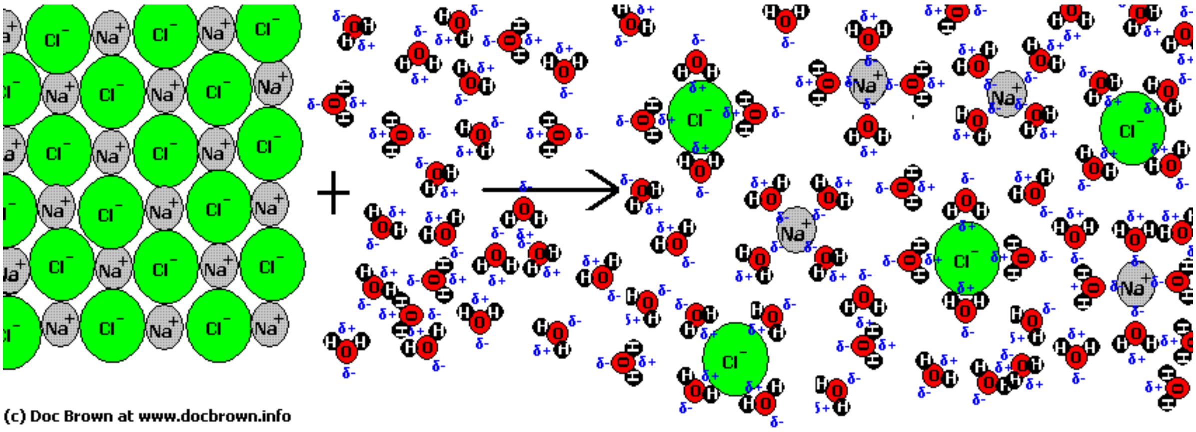
A Brass, a substitutional alloy



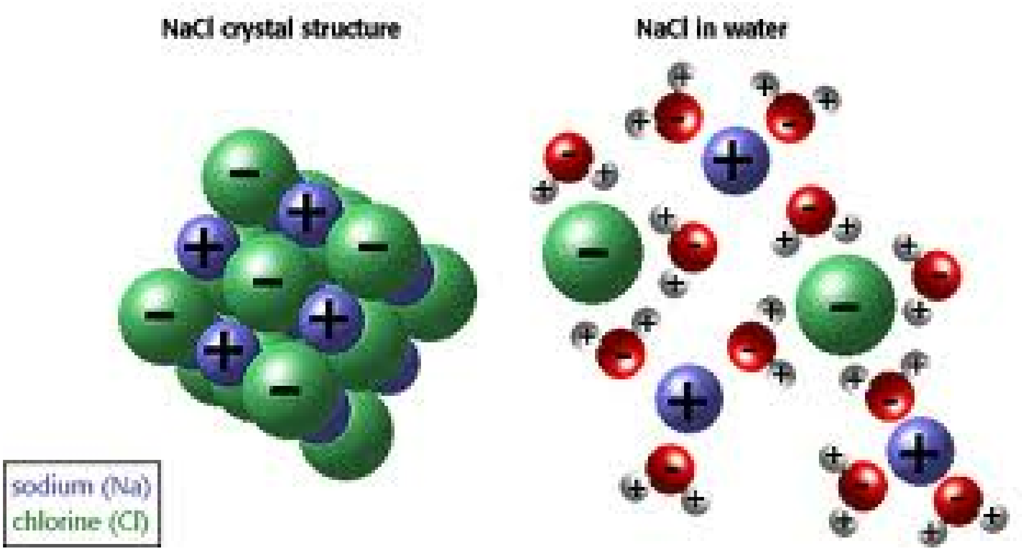
B Carbon steel, an interstitial alloy

What actually happens when sodium chloride "dissolves" in water?

Water molecules (solvent) collide with salt crystals (solute), which are in a **crystal lattice**, knocking Na^+ and Cl^- ions off the crystal and into the water.
Eventually, the crystal is gone and all the ions are dispersed among the water molecules.

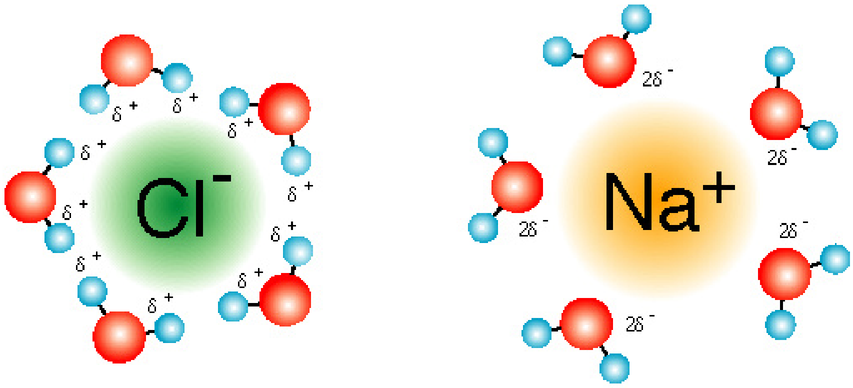


http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/molvie1.swf http://preparatorychemistry.com/Bishop_Solubility_frames.htm



Dissociation Equation:

$NaCl_{(s)} \longrightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$



Slightly positive hydrogen are attracted to chlorine anions

Slightly negative oxygen are attracted to sodium cations

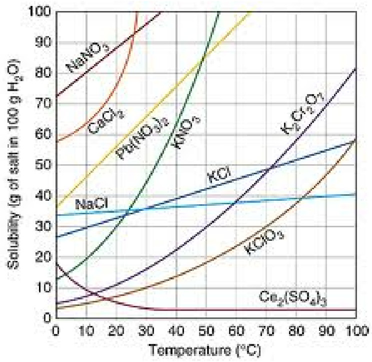
Because the crystal has been dispersed into individual ions, you can no longer see the salt in the water, and the result is a homogeneous mixture - a solution!

The ions are surrounded by water molecules (called hydration shells) as seen above.

http://www.northland.cc.mn.us/biology/Biology1111/animations/dissolve.html

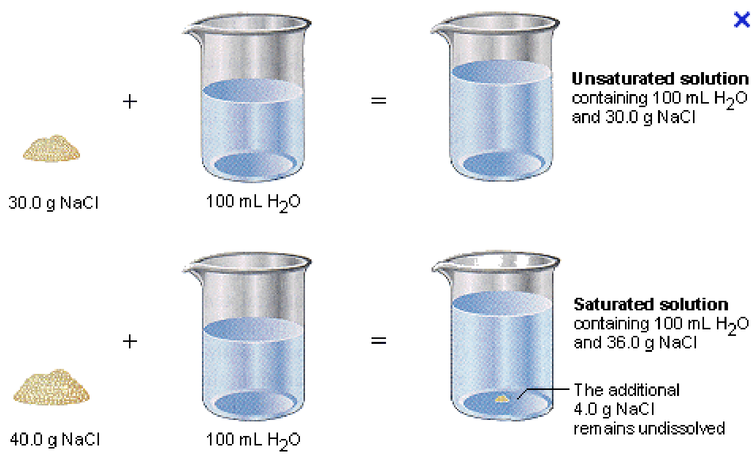
Temperature and Solutions

- Higher temperature makes the water molecules move faster and collide with the crystal lattice harder and more often
- Thus, a higher temperature speeds up dissolving.
- A higher temperature also increases the amount of solute that will dissolve.



Saturated Solutions

- when a solution is saturated, it has the maximum amount of solute possible for a given amount of solvent
- saturated solutions have excess solute remaining on the bottom of the container as no more can dissolve

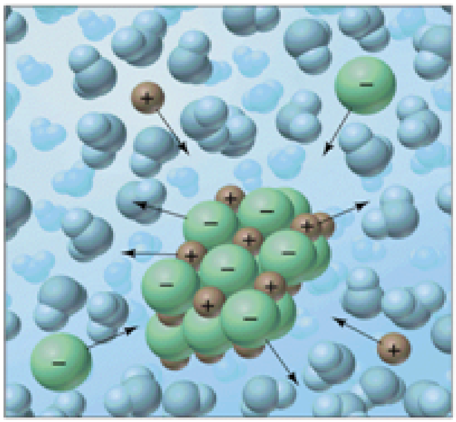


- Two processes are occurring in saturated solutions: solute crystals are dissolving AND the ions already dissolved in solution are "recrystallizing". These two processes happen at the same rate.
- Dissolving: $NaCl_{(s)} \Rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$
- Recrystallizing: $Na^{+}_{(aq)} + Cl^{-}_{(aq)} \Rightarrow NaCl_{(s)}$
- These two processes are opposites, and since they occur at the same rate, no net change results!

$$NaCl_{(s)} \rightleftharpoons Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

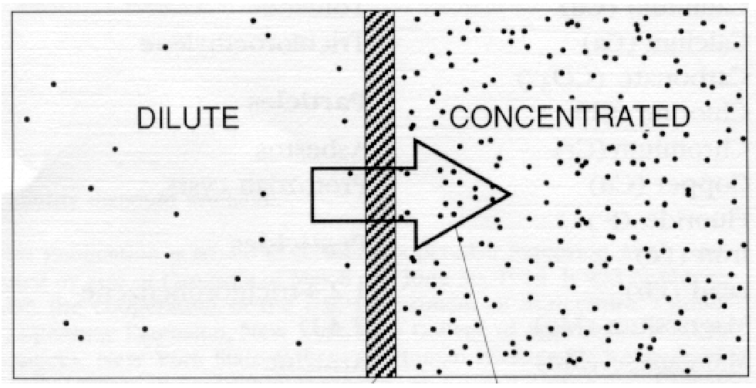
http://www.wwnorton.com/college/chemistry/gilbert2/tutorials/interface.asp?chapter=chapter_04&folder=saturated_solutions

http://www.dlt.ncssm.edu/core/Chapter14-Gas_Phase-Solubility-Complex_Ion_Equilibria/Chapter14-Animations/Solubility_of_AgCl.html



Concentration

- The concentration of a solution is the amount of solute in a given volume of solution
- If there is a lot of solute in a solution, the solution is **concentrated**. If there is a little solute in a solution, the solution is **dilute**.
- the amount of solute is usually measured in grams or moles
- the volume of solvent is usually measured in L or mL



If you have ever made Kool-Aid®, Gatorade®, orange juice concentrate, instant iced tea or any other solution mix you know how important it is to combine the right amount of mix with the right amount of water.

Too little mix and the drink is too weak - too much mix and the drink if too strong.
If you want it to taste just right you must follow the directions.

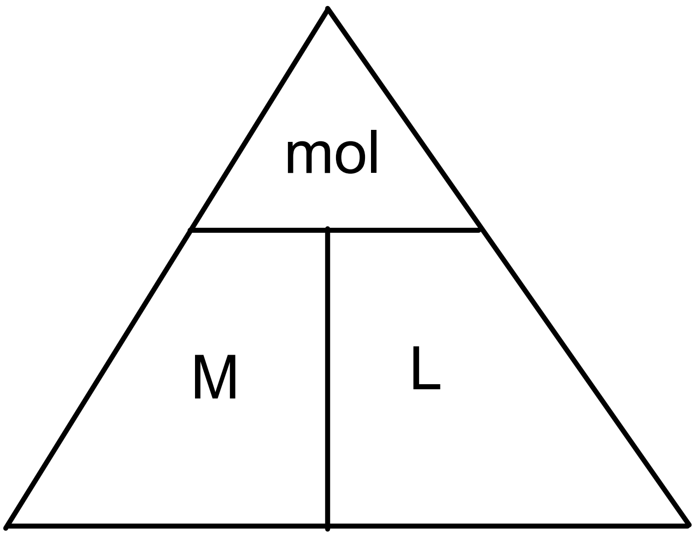
The mixture directions tell you the amount of **solute**
to put in the **solvent**
so that you come out with a solution that taste the way you expect it to taste.

Concentration can be most commonly measured in units of g/mL, g/L, mol/L

The most common unit of concentration is Molarity.

$$\text{Molarity} = \frac{\text{Moles of Solute}}{\text{Litres of Solution}}$$

Therefore, the units of Molarity are mol/L



The concentration of a substance can be described using square brackets around the formula. For example, [NaCl] means 'the concentration of NaCl'

Examples…

- Find the molarity of a solution made by 0.352mol of KCl in 2.33L of solution.
- M = mol solute/L solution
- M = 0.352mol/2.33L
- [KCl] = 0.151M
- There is 0.151 moles of KCl per litre of solution!

- How many moles of LiBr in 3.00L of a 0.5450M solution?
- mol = M x L
- mol = 0.5450M x 3.00L
- mol = 1.64 mol

How is Molarity used in the Lab?

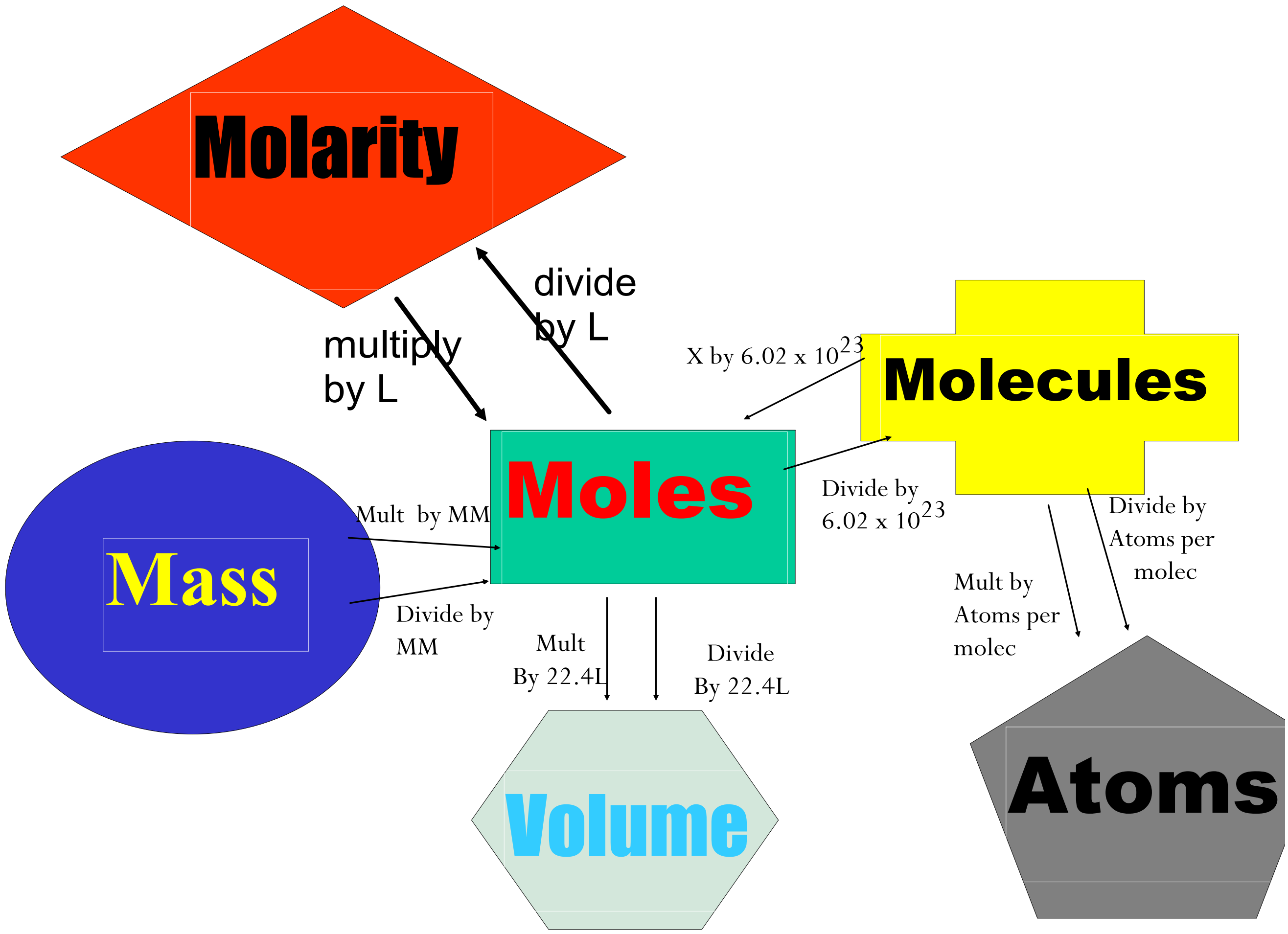
- What is [$CuSO_4$] if 25.0 g of copper (II) sulphate is dissolved in enough water to make 500.0 mL?
- Step 1: grams $CuSO_4$ to moles $CuSO_4$
- Step 2: moles $CuSO_4$ / volume (L) = Molarity

$$\frac{25.0\text{g}}{} \left| \frac{1 \text{ mole}}{159.7 \text{ g}} \right. = 0.157 \text{ moles}$$

$$\text{Molarity} = \frac{0.157\text{moles}}{0.500 \text{ L}} = 0.314 \text{ M}$$

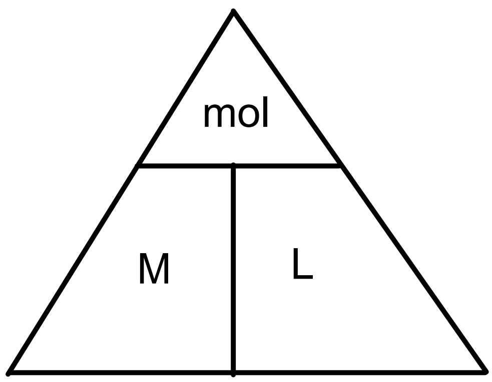
[$CuSO_4$] = 0.314M

MOLE MAP


Molarity
divide by L
multiply by L
X by 6.02 x 10^23
Molecules
Moles
Mass
Mult by MM
Divide by MM
Divide by 6.02 x 10^23
Divide by Atoms per molec
Mult by Atoms per molec
Mult By 22.4L
Divide By 22.4L
Volume
Atoms

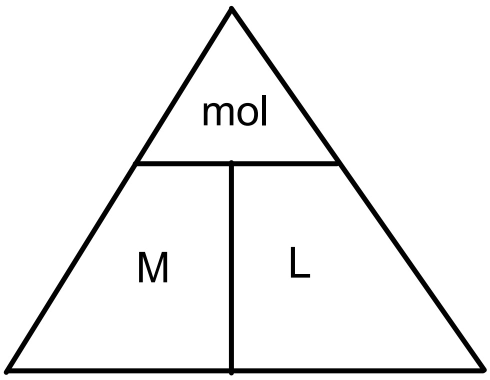
HOMEWORK:

- Read Hebden p. 96, 97, 98
- Molarity Worksheet #1-5
- (use Molarity triangle to help)



MOLARITY REVIEW

$$\text{Molarity} = \frac{\text{Moles of Solute}}{\text{Litres of Solution}}$$



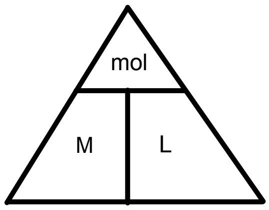
Concentration

- A measure of the amount of solute dissolved in a certain amount of solution.
- A concentrated solution has a large amount of solute.
- A dilute solution has a small amount of solute
- Concentration = $\dfrac{\text{amount of solute}}{\text{volume of solution}}$
- Sometimes g/L or g/mL or kg/L
- most often M (or mol/L)

Vocabulary Review

- Solubility- The amount of solute that will dissolve in a solution - common units are molarity (M = mol/L)
- Saturated solution- Contains the maximum amount of solid dissolved in a given volume of solution
- Unsaturated solution- Can still dissolve more solute

Molarity



- The number of moles of solute in 1 Litre of the solution.
- M = moles of solute/Litre of solution
- What is the molarity of a solution with 2.0 moles of NaCl in 4.0 Liters of solution?
- 0.50 M
-
- What is the molarity of a solution with 0.30 moles dissolved in 250 mL of solution?
- 1.2 M

Making solutions

How many moles of NaCl are needed to make 6.0 L of a 0.75 M NaCl solution?

- Moles = Molarity x Volume
- = 0.75 M x 6.0 L
- = 4.5 mol

Making solutions

- How many grams of $CaCl_2$ are needed to make 625 mL of a 2.00 M solution?
- Moles = M x L = 2.00 M x 0.625 L = 1.25 mol
- Mass $CaCl_2$ = $\frac{1.25\ \text{mol}}{} \Big| \frac{111.0\ \text{g}}{1\ \text{mol}}$
- = 139 g $CaCl_2$

http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/crm3s1_2.swf

Making solutions

- 10.3 g of NaCl is dissolved in 250 mL of water. What is [NaCl]?
- $\dfrac{10.3 \text{ g of NaCl} \mid 1 \text{ mole}}{\mid 58.5 \text{ g}} = 0.176 \text{ mol}$
-

$$M = \frac{0.176 \text{ mol}}{0.25 \text{ L}} = 0.70 \text{ M}$$

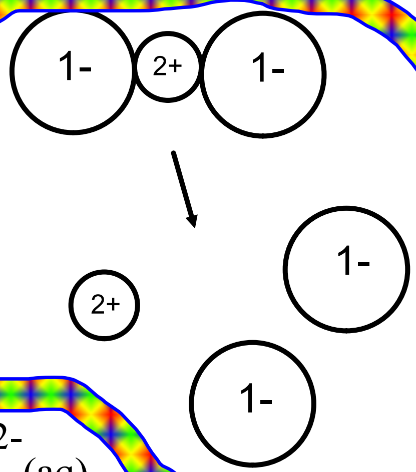
2. Dissolving

- A quick reminder that a salt is an ionic compound made up of a metal cation and a non-metal anion
- Salts dissolve in a process called *"dissociation"*
This means that the compound not only dissolves, but it 'dissociates' into ions.
- e.g. $KCl_{(s)} \Rightarrow K^{+}_{(aq)} + Cl^{-}_{(aq)}$

- Every salt dissolves in water to some extent. The amount of salt required to saturate the solution depends on the type of salt. Some salts can have a very high molarity before they become saturated, and some become saturated at very low molarities.
- If a salt has a saturated solution concentration greater than 0.1M, it is said to be *soluble*.
- If less than 0.1M, it is said to be *low solubility*

Dissociation Equations for Different Salts

- $LiF_{(s)} \Rightarrow Li^{+}_{(aq)} + F^{-}_{(aq)}$
- $MgI_{2(s)} \Rightarrow Mg^{2+}_{(aq)} + 2I^{-}_{(aq)}$
- (note: 2 iodide ions!)



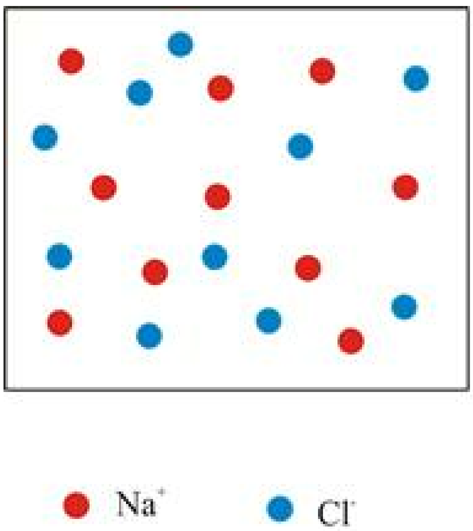
- $Na_2SO_{4(s)} \Rightarrow 2Na^{+}_{(aq)} + SO_4^{2-}{}_{(aq)}$
- $(NH_4)_2CO_3 \Rightarrow 2NH_4^{+}{}_{(aq)} + CO_3^{2-}{}_{(aq)}$
- Remember: Polyatomic compounds are also ionic in nature!

http://www.dlt.ncssm.edu/core/Chapter5-Moles-Molarity-Reaction_Types/Chapter5-Animations/Dissolving_NaCl-Electrolyte.html

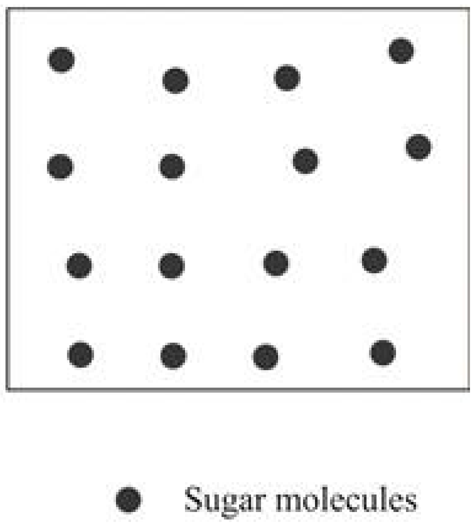
Covalent Compounds

- Covalent molecules have partial charges (if polar) or no charges (nonpolar)
- If a covalent compound can dissolve in water, it must be polar, and it dissolves as a whole molecule (it doesn't dissociate into ions).
- e.g. Sugar $C_{12}H_{22}O_{11(s)} \Rightarrow C_{12}H_{22}O_{11(aq)}$

http://www.dlt.ncssm.edu/core/Chapter5-Moles-Molarity-Reaction_Types/Chapter5-Animations/Dissolving_Sugar_Non-electrolyte.html



Salt in water

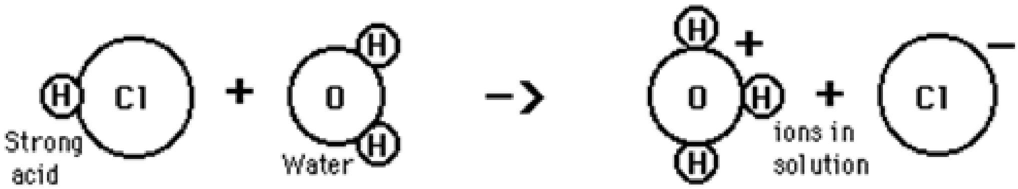


Sugar in water

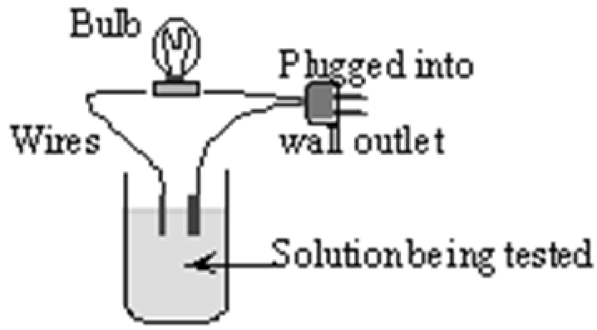
Acids

- Acids dissolve by reacting with water to form ions (see p. 74 for all acids)
- Strong partial negative dipole on O from H_2O removes H^+ from acid
- E.g. Hydrochloric acid

- $HCl_{(s)} + H_2O_{(l)} \Rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$



http://www.dlt.ncssm.edu/core/Chapter16-Acid-Base_Equilibria/Chapter16-Animations/StrongAcidIonization.html



Conductivity of Solutions

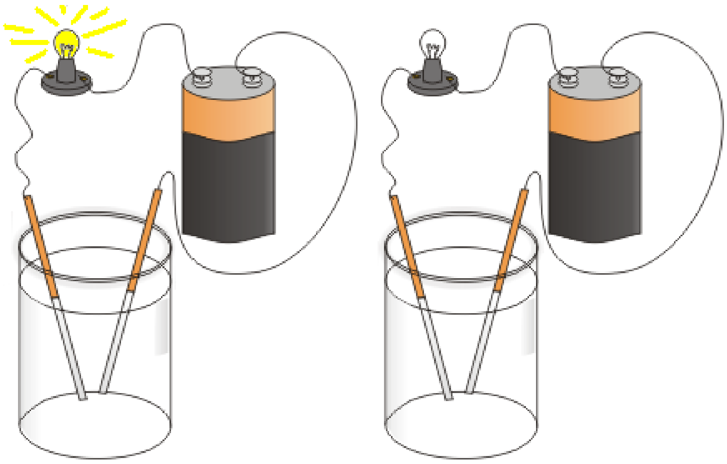
The Ability to Transmit Electricity

Conductivity

- conductivity is dependant on charged particles
- Ionic solutions conduct very well because charged particles (ions) are present
- Acid solutions conduct very well because ions are present
- Covalent solutions do not conduct electricity because no charges are present (entire neutral molecules are dissolved)

http://cwx.prenhall.com/petrucci/medialib/media_portfolio/05.html

- The more ions in the solution, the greater the conductivity
- Therefore concentrated solutions conduct better than dilute solutions

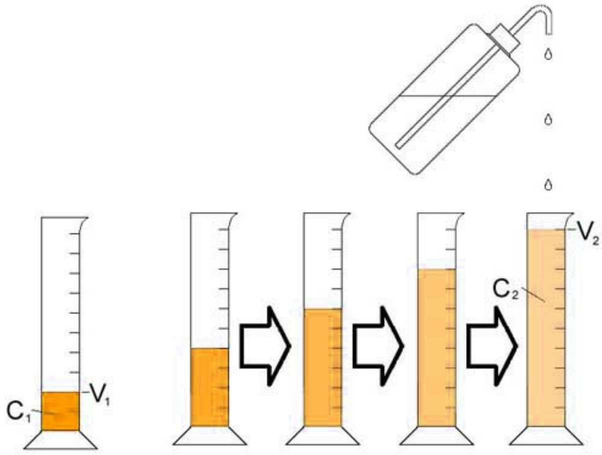


HOMEWORK:

- Molarity Worksheet #6-8
- Dissociation Equations/Ion Concentrations Worksheet - Part 1 ONLY
- Hebden p.198 #6

3. Dilution

(Adding water to a solution)

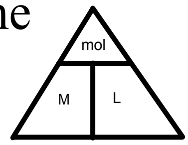


When solvent is added to dilute a solution, the number of moles of <u>solute</u> remains unchanged!

http://www.wwnorton.com/college/chemistry/gilbert2/tutorials/interface.asp?chapter=chapter_04&folder=dilutions

Moles solute before dilution = Moles solute after dilution

Since **moles of solute = M x V** (think of the molarity triangle), the dilution process can be described by the following equation:



mol solute before = mol solute after

$$(M_{initial})(V_{initial}) = (M_{final})(V_{final})$$

Initial is before the dilution has taken place, and Final is after!

Example

- **2.0 L of a 0.88 M solution is diluted to 3.8 L. What is the new molarity?**

$$M_i V_i = M_f V_f$$

$$M_f = \frac{M_i V_i}{V_f} = \frac{0.88M \times 2.0L}{3.8L}$$

$$= 0.46M$$

Notice that when diluting, the molarity decreases (0.88M to 0.46M).

Practice

- You have 150 mL of 6.0 M HCl. What volume of 1.3 M HCl can you make? $M_i V_i = M_f V_f$

$$V_f = \frac{M_i V_i}{M_f} = \frac{(6.0\,M)(0.150\,L)}{1.3M}$$

$$= 0.69 \text{ L}$$

$$= 690 \text{ mL}$$

Practice

- You need 450 mL of 0.15 M NaOH. All you have available is a 2.0 M stock solution of NaOH. How do you make the required solution?

$(2.0M)(V_i) = (0.15M)(0.45L)$

$V_i = 33.75mL = 34mL$

Measure out 34mL of the 2.0M NaOH and add water until you reach 450mL.

- You have 2.00L of a 0.500M solution of HCl. How much water must you add to dilute the solution to 0.315M?

$(0.500M)(2.00L) = (0.315M)(V_f)$

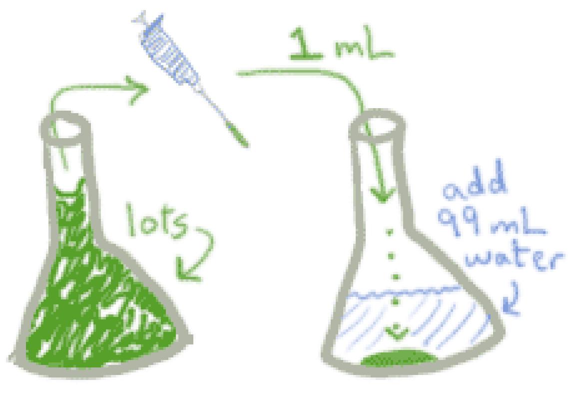
$V_f = 3.17L$

Volume added = 3.17 - 2.00 = 1.17L

http://highered.mcgraw-hill.com/sites/0072512644/student_view0/chapter4/animations_center.html#

HOMEWORK:

Dilution Worksheet



4. Calculating Ion Concentrations in Solution

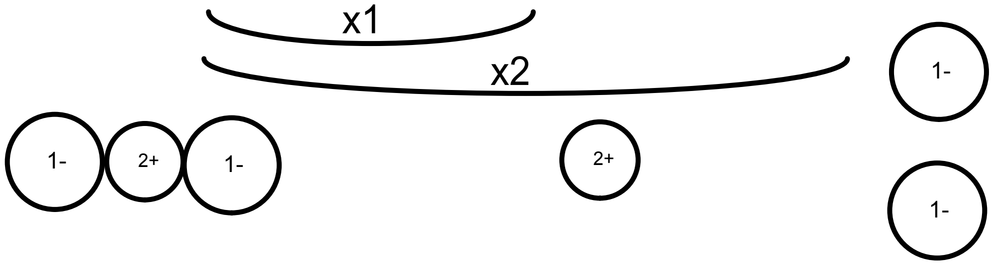
- Example of a DISSOCIATION EQUATION:

$$Sr(NO_3)_{2(s)} \longrightarrow Sr^{2+}_{(aq)} + 2NO_3^-{}_{(aq)}$$

Stoichiometric ratios apply to dissociation equations – can make mole-to-mole comparisons OR Molarity-to-Molarity comparisons based on coefficients.

$$Sr(NO_3)_{2(s)} \longrightarrow Sr^{2+}_{(aq)} + 2NO_3^-{}_{(aq)}$$

0.20M 0.20M 0.40M



- Determine the concentration of each ion present in 0.25 M K_2SO_4

$$K_2SO_{4(s)} \longrightarrow 2K^+_{(aq)} + SO_4^{2-}{}_{(aq)}$$

0.25M 0.50M 0.25M

- Determine the concentration of each ion present in 0.045M $AlBr_3$

$$AlBr_{3(s)} \longrightarrow Al^{3+}_{(aq)} + 3Br^-_{(aq)}$$

0.045M 0.045M 0.14M

- What volume (in L) of 0.25 M Na_3PO_4 would contain 0.80 mol of Na^+?

$$Na_3PO_{4(s)} \longrightarrow 3Na^+_{(aq)} + PO_4^{3-}{}_{(aq)}$$

0.25M 0.75M 0.25M

$$\frac{0.80mol}{0.75M} = 1.1L$$

- Determine the concentration of each ion in the following mixture:

  5.0 g KOH and 1.0 g $BaBr_2$ dissolved in 500.0 mL of solution.

$$\frac{5.0\text{g KOH}}{} \left| \frac{1 \text{ mol KOH}}{39.1\text{gKOH}} \right. = 0.\underline{12}79\text{mol KOH}$$

$$\frac{1.0\text{g BaBr}_2}{} \left| \frac{1 \text{ mol BaBr}_2}{297.1\text{g BaBr}_2} \right. = 0.00\underline{33}66 \text{ mol BaBr}_2$$

$$[KOH] = \frac{0.\underline{12}79\text{mol}}{0.5000\text{L}} = 0.\underline{25}6\text{M} \qquad [BaBr_2] = \frac{0.00\underline{33}66\text{mol}}{0.5000\text{L}} = 0.00\underline{67}32\text{M}$$

$KOH \longrightarrow$	$K^+_{(aq)}$ +	$OH^-_{(aq)}$	$BaBr_2 \longrightarrow$	Ba^{2+}	+	$2Br^-$
0.$\underline{25}$6M	0.26M	0.26M	0.00$\underline{67}$32M	0.0067M		0.013M

$[K^+]$ = 0.26M $[OH^-]$ = 0.26M $[Ba^{2+}]$ = 0.0067M $[Br^-]$ = 0.013M

- What is the concentration of each type of ion in a solution made by mixing 50.0 mL of 0.240 M $AlBr_3$ and 25.0 mL of 0.300 M $CaBr_2$?

$M_iV_i = M_fV_f$ V_f = 75.0mL = 0.0750L

$$[AlBr_3]_f = \frac{(0.240M)(0.0500L)}{(0.0750L)} = 0.160M$$

$$[CaBr_2]_f = \frac{(0.300M)(0.0250L)}{(0.0750L)} = 0.100M$$

$$\underset{0.160M}{AlBr_3} \longrightarrow \underset{0.160M}{Al^{3+}_{(aq)}} + \underset{0.480M}{3Br^-_{(aq)}}$$

$$\underset{0.100M}{CaBr_2} \longrightarrow \underset{0.100M}{Ca^{2+}_{(aq)}} + \underset{0.200M}{2Br^-_{(aq)}}$$

$[Al^{3+}]$ = 0.160M $[Ca^{2+}]$ = 0.100M $[Br^-]$ = 0.480M + 0.200M = 0.680M

HOMEWORK:

- Dissociation Equations/Ion Concentrations Worksheet - Part 2
- Hebden p. 212 #31, 35, 37, 38

5. Precipitation

No, not that kind!

Recrystallization

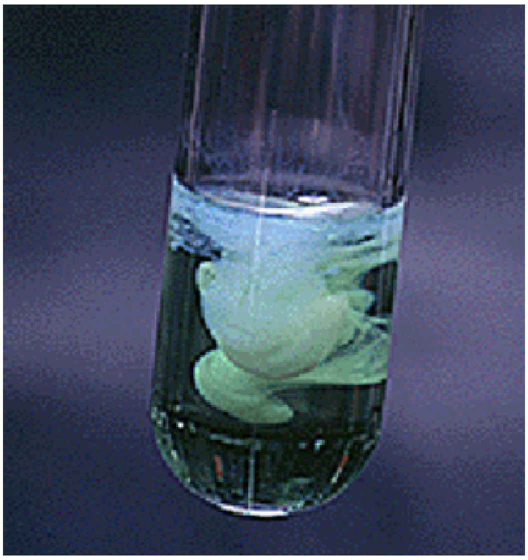
(formation of a solid)

Remember that EVERY salt dissociates to some extent in water. Some salts dissociate a great amount and have a high molarity at saturation, while others become saturated at a very low molarity.

A 'soluble' salt has a saturation molarity greater than 0.10M, whereas a 'low solubility' salt becomes saturated at a molarity lower than 0.10M.

- The Solubility Table allows us to determine which salts are 'soluble' and which are 'low solubility'

See Handout - Solubility Table



Solubility of Common Compounds in Water

The term soluble here means > 0.1 mol/L at 25°C.

Negative Ions (Anions)	Positive Ions (Cations)	Solubility of Compounds
All	Alkali ions: Li^+, Na^+, K^+, Rb^+, Cs^+, Fr^+	Soluble
All	Hydrogen ion: H^+	Soluble
All	Ammonium ion: NH_4^+	Soluble
Nitrate, NO_3^-	All	Soluble
Chloride, Cl^- or Bromide, Br^- or Iodide, I^-	All others	Soluble
	Ag^+, Pb^{2+}, Cu^+	Low Solubility
Sulphate, SO_4^{2-}	All others	Soluble
	Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}	Low Solubility
Sulphide, S^{2-}	Alkali ions, H^+, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}	Soluble
	All others	Low Solubility
Hydroxide, OH^-	Alkali ions, H^+, NH_4^+, Sr^{2+}	Soluble
	All others	Low Solubility
Phosphate, PO_4^{3-} or Carbonate, CO_3^{2-} or Sulphite, SO_3^{2-}	Alkali ions, H^+, NH_4^+	Soluble
	All others	Low Solubility

SOLUBILITY OF COMMON COMPOUNDS IN WATER

The term soluble here means > 0.1 mol/L at 25°C.

Negative Ions (Anions)	**Positive Ions (Cations)**	**Solubility of Compounds**
All	Alkali ions: Li^+, Na^+, K^+, Rb^+, Cs^+, Fr^+	Soluble
All	Hydrogen ion: H^+	Soluble
All	Ammonium ion: NH_4^+	Soluble

	Cations	
Nitrate, NO_3^-	All	Soluble

Chloride, Cl^- or Bromide, Br^- or Iodide, I^-	All others	Soluble
	Ag^+, Pb^{2+}, Cu^+	Low Solubility
Sulphate, SO_4^{2-}	All others	Soluble
	Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}	Low Solubility
Sulphide, S^{2-}	Alkali ions, H^+, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}	Soluble
	All others	Low Solubility

Hydroxide, OH^-	Alkali ions, H^+, NH_4^+, Sr^{2+} ?	Soluble
	All others	Low Solubility
Phosphate, PO_4^{3-} or Carbonate, CO_3^{2-} or Sulphite, SO_3^{2-}	Alkali ions, H^+, NH_4^+	Soluble
	All others	Low Solubility

Use your table to predict whether the following salts are soluble (S) or low solubility (LS) and whether they form a precipitate (ppt) in water

- 1. Sodium hydroxide S
- 2. Ammonium acetate S
- 3. Calcium sulphate LS (ppt)
- 4. Lead II chloride LS (ppt)
- 5. Potassium chloride S
- 6. Calcium bromide S
- 7. Potassium carbonate S
- 8. Aluminum sulphate S
- 9. Copper II chloride S
- 10. Copper I chloride LS (ppt)

Suppose you wanted to make a saturated solution of PbI_2. One way you could do this is to dissolve (dissociate) PbI_2 in water (making $Pb^{2+}_{(aq)}$ and $I^-_{(aq)}$) until no more will dissolve and you have excess $PbI_{2(s)}$ on the bottom.

http://www.wwnorton.com/college/chemistry/gilbert2/tutorials/interface.asp?chapter=chapter_04&folder=saturated_solutions

Another way is to mix one solution that has $Pb^{2+}_{(aq)}$ ions to another solution that has $I^-_{(aq)}$ ions.

Let`s suppose you decided to mix equal volumes of two soluble salt solutions together, such as 0.20M $KI_{(aq)}$ with 0.20M $Pb(NO_3)_{2(aq)}$

$KI_{(aq)}$ is actually $K^+_{(aq)}$ and $I^-_{(aq)}$

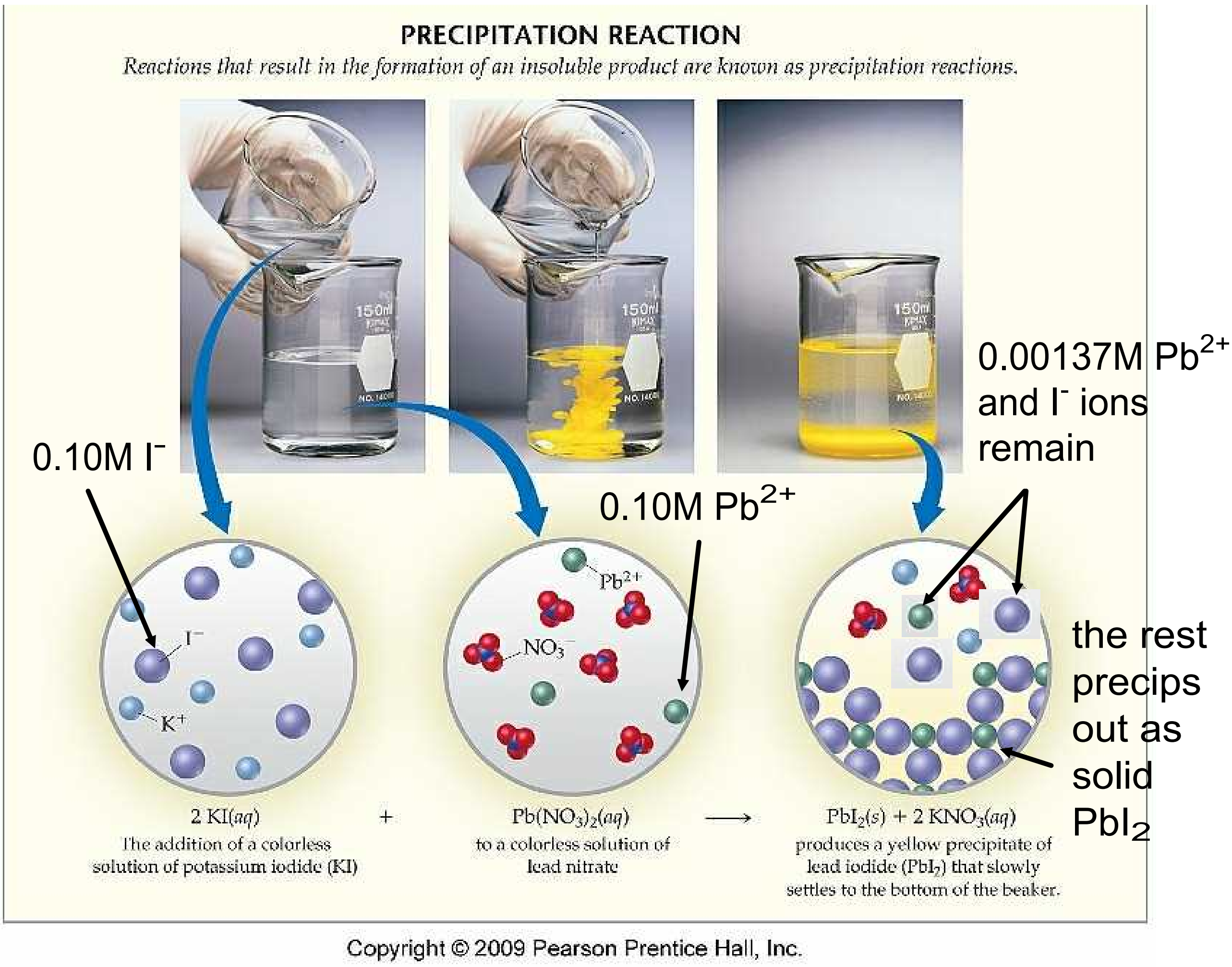
$Pb(NO_3)_{2(aq)}$ is actually $Pb^{2+}_{(aq)}$ and $NO_3^-{}_{(aq)}$

By mixing, you`ve introduced Pb^{2+} to I^- and also K^+ to NO_3^-.

If either of these combinations are `low solubility` together, they will be 'oversaturated' and precipitate out of solution (form a solid). This will be the case for Pb^{2+} and I^- and they will form the precipitate $PbI_{2(s)}$, thereby creating a saturated solution of PbI_2.

http://www.dlt.ncssm.edu/core/Chapter5-Moles-Molarity-Reaction_Types/Chapter5-Animations/DoubleDisp_Reaction-Precipitation.html

- When two soluble ionic solutions that have molarities greater than 0.10M are mixed together and at least one new combination of cation and anion have low solubility, a ***precipitate*** forms
- On the next slide, Pb^{2+} and I^{-} are low solubility together, so when mixed together, they will precipitate out of solution. The saturation molarity for PbI_2 is 000137M, so solid PbI_2 will precipitate out until that molarity remains in solution as aqeuous ions


PRECIPITATION REACTION
Reactions that result in the formation of an insoluble product are known as precipitation reactions.
0.10M I⁻
0.10M Pb2+
0.00137M Pb2+ and I⁻ ions remain
the rest precips out as solid PbI2
I⁻
K⁺
Pb2+
NO3⁻
2 KI(aq) + Pb(NO3)2(aq) ⟶ PbI2(s) + 2 KNO3(aq)
The addition of a colorless solution of potassium iodide (KI)
to a colorless solution of lead nitrate
produces a yellow precipitate of lead iodide (PbI2) that slowly settles to the bottom of the beaker.
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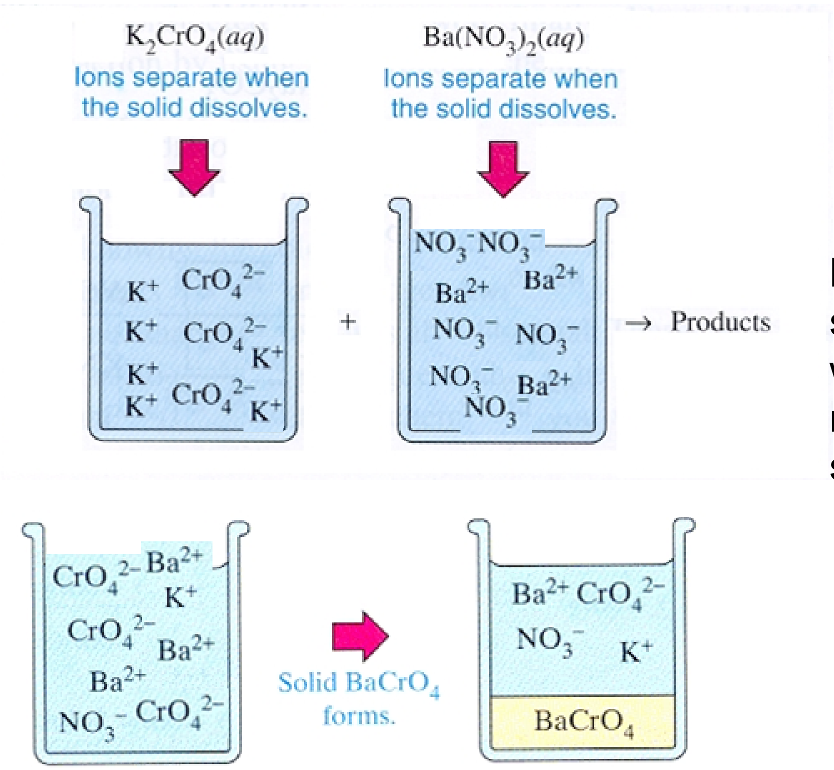


Ba^{2+} and CrO_4^{2-} are low solubility together, so when mixed together, most will precipitate out of solution as a solid.

*Note: CrO_4^{2-} is not on your table

When mixing two ionic solutions three outcomes can result:

- No precipitate forms (all combinations of cation & anion are soluble)

- One precipitate forms (one combination of cation & anion have low solubility)

- Two precipitates form (both combinations of cation and anion have low solubility) http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/crm3s2_3.swf

What would result if 1.0M solutions of $Al_2(SO_4)_{3(aq)}$ and $Sr(OH)_{2(aq)}$ were mixed together?

Two precipitates would form:
$Al(OH)_{3(s)}$ and $SrSO_{4(s)}$

HOMEWORK:

Using the Solubility Table Worksheet

6. Formula, Complete, & Net Ionic Equations

http://preparatorychemistry.com/Bishop_Solubility_frames.htm

We can describe the mixture of two ionic solutions in three ways:

- Formula Equations
- Complete Ionic Equations
- Net Ionic Equations

Steps for Formula Equations

- Complete the double replacement reaction
- Balance each equation
- Look up each product on the solubility table
- If it is soluble, write (aq) after the formula
- If it is low solubility write (s) after the formula

Formula Equations
- no ions are written!

1. $NaCl_{(aq)} + KNO_{3(aq)} \Rightarrow NaNO_{3(aq)} + KCl_{(aq)}$
2. $Na_2CO_{3(aq)} + BaCl_{2(aq)} \Rightarrow 2NaCl_{(aq)} + BaCO_{3(s)}$
3. $Sr(OH)_{2(aq)} + MgSO_{4(aq)} \Rightarrow Mg(OH)_{2(s)} + SrSO_{4(s)}$

*Note: look at equation #1. Even though $NaNO_{3(aq)}$ is written as a product, the $Na^{+}_{(aq)}$ ion and $NO_3^{-}{}_{(aq)}$ ion do not actually come together. They stay separated in solution. Same idea for the $KCl_{(aq)}$ and the $NaCl_{(aq)}$ in equation #2.

Complete Ionic Equations
- all ions except ppts
-the most accurate equation

1. $Na^{+}_{(aq)} + Cl^{-}_{(aq)} + K^{+}_{(aq)} + NO_3^{-}{}_{(aq)} \Rightarrow$
   $Na^{+}_{(aq)} + NO_3^{-}{}_{(aq)} + K^{+}_{(aq)} + Cl^{-}_{(aq)}$

2. $2Na^{+}_{(aq)} + CO_3^{2-}{}_{(aq)} + Ba^{2+}_{(aq)} + 2Cl^{-}_{(aq)} \Rightarrow$
   $2Na^{+}_{(aq)} + 2Cl^{-}_{(aq)} + BaCO_{3(s)}$

3. $Sr^{2+}_{(aq)} + 2OH^{-}_{(aq)} + Mg^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \Rightarrow$
   $SrSO_{4\ (s)} + Mg(OH)_{2(s)}$

- Ions that do not change from reactant to product are not technically part of a chemical reaction.
- They are called ***spectator ions***.

- For a ***net ionic equation*** simply remove the spectator ions from the equation, as a **<u>net</u>** equation only includes the substances that have reacted.

Net Ionic Equations
- only ions that change
(make a ppt)

1. No net ionic equation (all spectator ions)

2. $Ba^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)} \Rightarrow BaCO_{3(s)}$
spectator ions were $Na^{+}_{(aq)}$ and $Cl^{-}_{(aq)}$

3. $Sr^{2+}_{(aq)} + 2OH^{-}_{(aq)} + Mg^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \Rightarrow$
$SrSO_{4\,(s)} + Mg(OH)_{2(s)}$
- same as the complete ionic equation
- no spectator ions

Write the formula, complete ionic, and net ionic equation for a mixture of $NaCl_{(aq)}$ and $Pb(NO_3)_{2(aq)}$

F: $2NaCl_{(aq)} + Pb(NO_3)_{2(aq)} \longrightarrow 2NaNO_{3(aq)} + PbCl_{2(s)}$

C: $2Na^+_{(aq)} + 2Cl^-_{(aq)} + Pb^{2+}_{(aq)} + 2NO_3^-{}_{(aq)} \longrightarrow 2Na^+_{(aq)} + 2NO_3^-{}_{(aq)} + PbCl_{2(s)}$

N: $Pb^{2+}_{(aq)} + 2Cl^-_{(aq)} \longrightarrow PbCl_{2(s)}$

HOMEWORK:

F, C, N Equations Worksheet

7. QUALITATIVE ANALYSIS

Finding out WHAT you have, not HOW MUCH

- **Qualitative Analysis** describes a procedure to determine which ions are present in a solution
- Suppose we want to test if an ion is in a solution
- The idea is to use the solubility table to choose an oppositely charged ion that is low solubility with the ion in question. If a precipitate forms, the ion in question is present. If not, that ion was not present to start with

- If there is more than one ion present in the solution, they should be tested one ion at a time (i.e. choose an oppositely charged ion that is low solubility with only ONE of the ions in question)

Steps for a mixture of CATIONS :

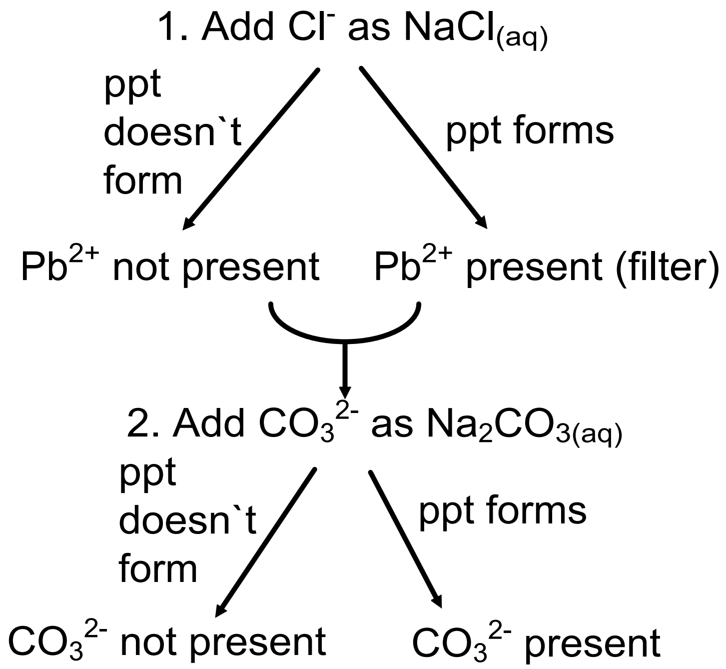
- Choose an ***anion*** that can potentially form a precipitate with **only one** of the cations.
- Filter out precipitate, then choose another anion to potentially precipitate the next cation.
- Filter out precipitate. Repeat if necessary until all cations are isolated in separate precipitates.

Steps for a mixture of ANIONS

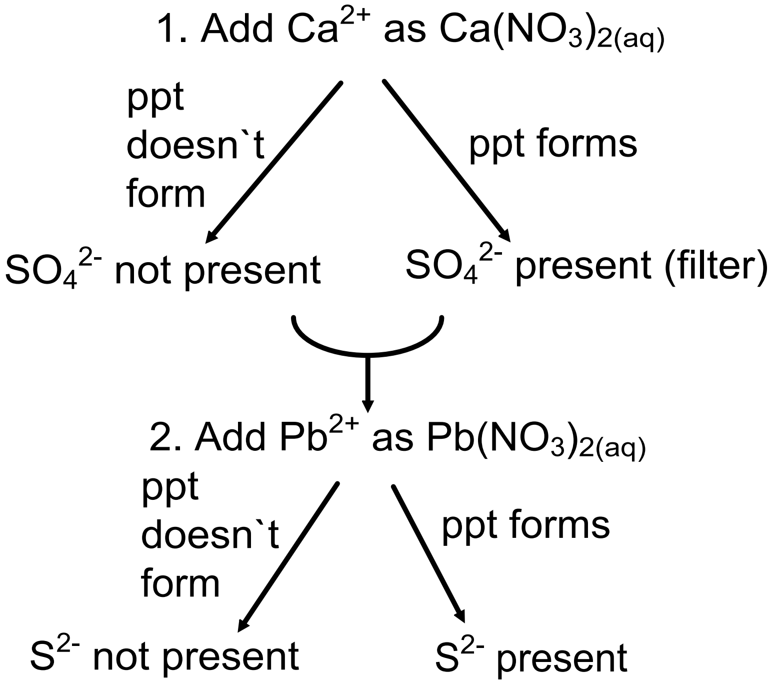
- Same procedure, but you will choose appropriate ***cations*** for each step.

- Note that cations added to a solution are usually added with NO_3^- (since it's a spectator and will not interfere) and anions are usually added with Na^+ or K^+ since they are soluble with any anion

Example - Test whether the following ions are in solution: Sr^{2+} and Pb^{2+}



Example: Test whether the following anions are in solution: S^{2-} and SO_4^{2-}



<u>HOMEWORK:</u>

Separating Ions from Solution Worksheet