

Tuesday, September 11, 2012  
12:14 PM



## Solutions

Inserted from: <<file:///C:/Users/Jaymo/Desktop/Chem11/PDF/Solutions.pdf>>

## Solution Chemistry

### Nature of Solutions

- solutions are **homogeneous** mixtures
- substances in solution are different from their solid, liquid or gas forms
- there should be no observable segregation of component substances

looks the same

### Types of Solutions

#### 1. Gaseous solutions

**Ex.** Air - 78% N<sub>2</sub>, 21% O<sub>2</sub>, 1% Ar, 0.03% CO<sub>2</sub>

#### 2. Solid solutions (alloys)

**Ex.** Sterling silver (Cu + Ag), brass (Cu + Zn),  
Dental amalgam (Hg + Ag), steel (Fe + C)

#### 3. Liquid solutions:

a. liquid - liquid

**Ex.** vinegar

b. solid - liquid

**Ex.** salt water

c. gas - liquid

**Ex.** carbonated water in soda pop

- The substance present in the greater amount is the **solvent**
- The **solute** is the component of lesser abundance

## Polar versus Non-polar Solvents

### Molecular polarity

A dipole results when electrons are displaced to one side of the molecule

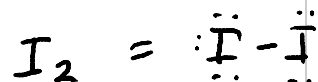
**Ex.** H<sub>2</sub>O - polar

CCl<sub>4</sub> - non-polar

### Polar versus Non-polar

**Polar** solvents tend to dissolve **polar** (and ionic) solutes

**Non-polar** solvents tend to dissolve **non-polar** solutes



### Water

Solutions containing water are called **aqueous (aq)**

Many important reactions (Ex. in living things) take place in aqueous solutions  
(Probably) the most common solvent

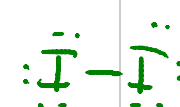
solvent

### Saturated and Unsaturated Solutions

- **Saturated** - solution in which no additional solute can dissolve
- **Unsaturated** - a solution in which additional solute can dissolve
- **Miscible** - when there is no apparent limit to the solubility of one liquid in another
- **Immiscible** - two liquids which do not dissolve in each other; solvation (interaction between solute and solvent particles) does not occur
- **Solubility** - is a measure of the amount of solute which can dissolve in a given amount of solvent
  - g/L, g/mL, g/g, mL/mL, are some units for this
  - adding a little at a time until saturated; or evaporating a saturated solution and measuring the mass of the solid left behind, are ways to determine solubility
  - Solubility Table defines soluble as  $> 0.1 \text{ mol/L @ } 25^\circ\text{C}$

Pol/Pol  
non-P/non-l  
← non-P/pol

Molarity



### Concentration of Solutions

Concentration describes the amount of solute in a given amount of solutions.

a) % solution      X g / 100 g of solution (W/W)  
                          X mL / 100 mL of solution (V/V)

b) Molarity    →     $\frac{\text{\# of moles of solute}}{\text{litres of solution made}}$



$$M = n/V \quad n = \# \text{ mols, } V = \text{volume (L)}$$
$$n = M \times V$$
$$V = n/M$$

"relative"

c) Dilute versus Concentrated

- dilute solutions contain less solute per litre of solution than concentrated solutions

**Ex.** Dilute HCl: 0.1M or 0.5M  
Concentrated HCl: 12M

Lab 16A read (w.s. provided)  
HW pg 194 # 1-4

### Dilution Calculations

Often solutions are made by diluting a more concentrated (stock) solution

$$n = M \times V$$

$$\text{Stock sol'n (conc.): } n_1 = M_1 \times V_1$$

$$\text{New sol'n (dilute): } n_2 = M_2 \times V_2$$

$$n_1 = n_2$$

So...

$$M_1 \times V_1 = M_2 \times V_2$$

(# moles does not change, only M does!)

$$M_i V_i = M_f V_f$$

**Ex.** If you have 500.0 mL of 0.40 M solution, what volume of a 0.10 M solution can you make? **How much water must be added?**

$$\begin{aligned} \# \text{ moles} &= 0.40 \text{ M} \times 0.5000 \text{ L} = 0.20 \text{ mol} \\ \text{Vol. solution to be made} &= 0.20 \text{ mol} / 0.10 \text{ M} = 2.0 \text{ L} \\ \text{Vol. H}_2\text{O to be added} &= 2.0 \text{ L} - 0.5000 \text{ L} = 1.5 \text{ L} \end{aligned}$$

OR

$$\begin{aligned} M_1 \times V_1 &= M_2 \times V_2 \\ 0.40 \text{ M} (0.5000 \text{ L}) &= (0.10 \text{ M}) (V_2) \\ V_2 &= \frac{(0.40 \text{ M}) (0.5000 \text{ L})}{0.10 \text{ M}} = \mathbf{2.0 \text{ L}} \end{aligned}$$

**Ex.** **What final concentration** results when 150 mL of 0.36 M solution of  $\text{MgSO}_4$  is added to 750 mL of water?

$$\begin{aligned} \text{moles MgSO}_4 &= 0.36 \text{ M} \times 0.150 \text{ L} = 0.054 \text{ mol} \\ \text{final vol.} &= 0.150 \text{ L} + 0.750 \text{ L} = 0.900 \text{ L} \\ \text{molarity} &= 0.054 \text{ mol} / 0.900 \text{ L} = \mathbf{0.060 \text{ M}} \end{aligned}$$

OR

$$\begin{aligned} (0.36 \text{ M}) (0.150 \text{ L}) &= (M_2) (0.900 \text{ L}) \\ \text{so } M_2 &= \frac{(0.36 \text{ M}) (0.150 \text{ L})}{0.900 \text{ L}} = \mathbf{0.060 \text{ M}} \end{aligned}$$

**Ex.** **What volume** of 0.50 M  $\text{CuCl}_2$  stock solution is required to make 2.0 L of 0.20 M  $\text{CuCl}_2$ ? *vi?* *Mi*

*v<sub>f</sub>* *M<sub>f</sub>* *now much H<sub>2</sub>O?*

mols  $\text{CuCl}_2$  needed = 2.0 L x 0.20 M = 0.40 mol

vol. =  $n/M = \frac{0.40 \text{ mol}}{0.50 \text{ M}}$

= 0.80 L of 0.50 M  $\text{CuCl}_2$

OR

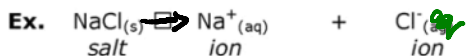
$0.50 \text{ M} (V_1) = (0.20 \text{ M})(2.0 \text{ L})$

so  $V_1 = \frac{(0.20 \text{ M})(2.0 \text{ L})}{0.50 \text{ M}} = 0.80 \text{ L}$

vol. of  $\text{H}_2\text{O}$  to be added = 2.0 L - 0.80 L = **1.2 L** *V<sub>f</sub> - V<sub>i</sub>*

### Aqueous solutions of ionic substances

Ionic substances dissolve in water to form ions



The concentration of ions can be different from the concentrations of the salts



$[\text{CaCl}_2] = \mathbf{0.10 \text{ M}}$

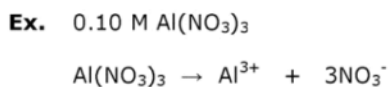
$[\text{Ca}^{2+}] = \mathbf{0.10 \text{ M}}$

$[\text{Cl}^-] = 2 \times 0.10 \text{ M} = \mathbf{0.20 \text{ M}}$

This is because every mole of  $\text{CaCl}_2$  dissolves to give one mole of  $\text{Ca}^{2+}$  ions and two moles of  $\text{Cl}^-$  ions.

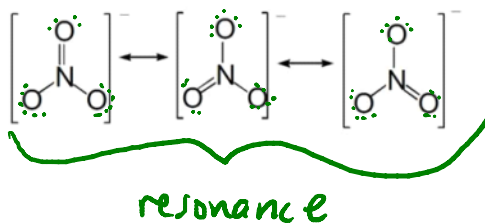
### Examples:

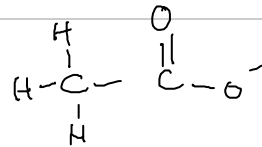
Write a dissociation equation for each of the following and give the concentration of the ions:



$[\text{Al}^{3+}] = \mathbf{0.10 \text{ M}}$

$[\text{NO}_3^-] = 3 \times 0.10 \text{ M} = \mathbf{0.30 \text{ M}}$





**Ex.** 0.35 M  $\text{Cu}(\text{CH}_3\text{COO})_2$



$[\text{Cu}^{2+}] = 0.35 \text{ M}$

$[\text{CH}_3\text{COO}^-] = 2(0.35 \text{ M}) = 0.70 \text{ M}$

**Ex.** 0.25 M  $\text{Na}_2\text{O}$

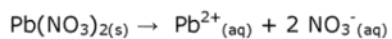


$[\text{Na}^+] = 2(0.25 \text{ M}) = 0.50 \text{ M}$

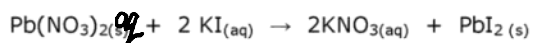
$[\text{O}^{2-}] = 0.25 \text{ M}$

### Equation types

*Dissociation Eq<sup>n</sup>:*

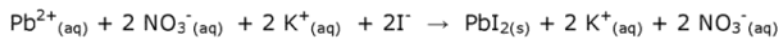


*Formula Eq<sup>n</sup>:*



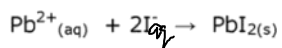
Simple-  
gr 10.

*Complete (Full) Ionic Eq<sup>n</sup>:*



*Net Ionic Eq<sup>n</sup>:*

Shows only the reaction species in a reaction in an aqueous environment (spectator ions are deleted)



## Qualitative Analysis

- Procedures by which a chemist decides what substances or ions are present in a sample
- Use the Solubility Table to select a reagent that gives a precipitate (ppt) with some substances but not with others
- If an anion is added to give a ppt, use an alkali metal salt of the anion (eg.  $\text{Na}^+$  or  $\text{K}^+$ ) - they will not ppt with other anions
- If adding a cation, use the nitrate salt of the cation
- Generally ppt the "easier to ppt" ion first

$\oplus$  ion       $\ominus$  ion

**Ex.** Devise a procedure for separating a mixture of  $\text{Ba}^{2+}$  and  $\text{Pb}^{2+}$  by precipitating (ppt) them individually.

- $\text{Pb}^{2+}$  ppts with  $\text{Cl}^-$ ,  $\text{Br}^-$ ,  $\text{I}^-$ , and  $\text{S}^{2-}$  - but  $\text{Ba}^{2+}$  is soluble
- Add a reagent containing  $\text{Cl}^-$ ,  $\text{Br}^-$ ,  $\text{I}^-$ , or  $\text{S}^{2-}$   
**Ex.**  $\text{NaCl}$ ,  $\text{KI}$ ,  $\text{LiBr}$
- $\text{Pb}^{2+}_{(\text{aq})} + 2 \text{Br}^{-}_{(\text{aq})} \rightarrow \text{PbBr}_{2(\text{s})}$
- filter off ppt; only  $\text{Ba}^{2+}$  left
- add a reagent containing  $\text{SO}_4^{2-}$ ,  $\text{CO}_3^{2-}$ ,  $\text{PO}_4^{3-}$ , or  $\text{SO}_3^{2-}$  ions to ppt  $\text{Ba}^{2+}$   
**Ex.**  $\text{Na}_3\text{PO}_4$
- $3 \text{Ba}^{2+}_{(\text{aq})} + 2 \text{PO}_4^{3-}_{(\text{aq})} \rightarrow \text{Ba}_3(\text{PO}_4)_{2(\text{s})}$

**Ex.** Devise a procedure to separate a mixture of  $\text{CO}_3^{2-}$  and  $\text{Cl}^-$

Anions chose cation  $\oplus$

- ppt  $\text{CO}_3^{2-}$  first using a cation other than alkali ions,  $\text{H}^+$  or  $\text{NH}_4^+$  (both are soluble) or  $\text{Ag}^+$ ,  $\text{Cu}^+$ ,  $\text{Pb}^{2+}$  (both will ppt)  
**Ex.**  $\text{Ca}^{2+}$  as  $\text{Ca}(\text{NO}_3)_2$
- $\text{Ca}^{2+}_{(\text{aq})} + \text{CO}_3^{2-}_{(\text{aq})} \rightarrow \text{CaCO}_{3(\text{s})}$
- Now ppt  $\text{Cl}^-$  using either  $\text{Ag}^+$ ,  $\text{Cu}^+$ ,  $\text{Pb}^{2+}$   
**Ex.**  $\text{Ag}^+$  as  $\text{AgNO}_3$
- $\text{Ag}^+_{(\text{aq})} + \text{Cl}^{-}_{(\text{aq})} \rightarrow \text{AgCl}_{(\text{s})}$

## Titration

An important analytical technique, titrations are used to determine the concentration of a substance in solution by comparing it to a solution of a known concentration; called a **standard**

The reaction between the unknown and the standard is allowed to proceed until mole ratios like those in the balanced chemical equation are reached.

This stoichiometric point (sometimes called the equivalence point or end point) is determined using some sort of chemical indicator

Often these reactions are used with acids/bases, for acid/base titrations the end point is reached when:

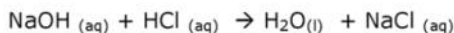
### Moles of OH<sup>-</sup> from the base = moles of H<sup>+</sup> from the acid

Titrations consist of several trials; the first is an estimate, usually done quickly to get an approximate endpoint, almost always the first trial will be over the endpoint (overshot) The average volume of the second and subsequent trials are used in calculations:

#### Steps to follow

- 1) Write the balanced equation (as always – you can't go wrong with this)
- 2) Calculate the moles of standard used ( $n = M \times V$ )
- 3) Determine the moles of unknown using the mole ratio
- 4) Calculate the unknown concentration ( $M = n/V$ )

**Ex.** If 10.0mL of HCl is titrated with 20.0mL of 0.40M NaOH, what is [HCl]?



Moles of standard (in this case NaOH) =  $0.40 \text{ M} \times 0.0200 \text{ L} = 8.0 \times 10^{-3}$

Moles of unknown (HCl) =

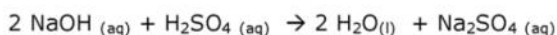
$$8.0 \times 10^{-3} \text{ mol NaOH} \left( \frac{1 \text{ mol HCl}}{1 \text{ mol NaOH}} \right) = 8.0 \times 10^{-3} \text{ mol HCl}$$

$$[\text{HCl}] = \left( \frac{8.0 \times 10^{-3} \text{ mol HCl}}{0.0100 \text{ L}} \right) = 0.80 \text{ M HCl}$$

- Some acids or bases produce two (or more) H<sup>+</sup> or OH<sup>-</sup> for each formula unit (H<sub>3</sub>PO<sub>4</sub> → 3 H<sup>+</sup> + PO<sub>4</sub><sup>3-</sup>)



**Ex.** An average volume of 24.90mL of 0.100 M NaOH was required to neutralize 15.00mL of H<sub>2</sub>SO<sub>4</sub>. What is the concentration of the acid?



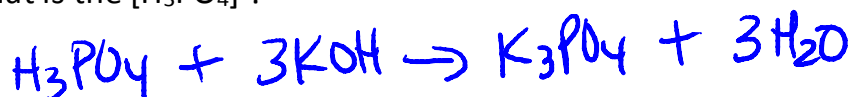
Moles of standard (in this case NaOH) = 0.100 M x 0.02490 L = 2.49 x 10<sup>-3</sup>

Moles of unknown (HCl) =

$$2.49 \times 10^{-3} \text{ mol NaOH} \left( \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} \right) = 1.25 \times 10^{-3} \text{ mol H}_2\text{SO}_4$$

$$[\text{HCl}] = \left( \frac{1.25 \times 10^{-3} \text{ mol H}_2\text{SO}_4}{0.01500 \text{ L}} \right) = 0.0833 \text{ M H}_2\text{SO}_4$$

Ex. 3 A 15.00mL sample of phosphoric acid is titrated to a phenolphthalein end point with 27.24 mL of a 0.502M KOH solution. What is the [H<sub>3</sub>PO<sub>4</sub>] ?



$$27.24 \text{ mL} \times 0.502 \text{ M KOH} = \text{moles KOH}$$

$$- \times \frac{1 \text{ H}_3\text{PO}_4}{3 \text{ KOH}} = \text{moles H}_3\text{PO}_4$$

$$- \div 15 \text{ mL} = [ ] \underline{\underline{0.307 \text{ M}}}$$

Hebden pg 212 # 30-38

Qual. Analysis H/O

LAB 16 D - prep (handout)