**CHM152LL Solution Chemistry Worksheet**

Many chemical reactions occur in solution. Solids are often dissolved in a solvent and mixed to produce a chemical reaction that would not occur if the solids themselves were directly mixed together. This worksheet will review the solution process, the different types of chemical reactions that occur in solution, molecular-level drawings of the reactions, and calculations commonly used in the laboratory.

**Solutions & Solubility Rules:** When a solute is dissolved in a solvent we call the mixture a solution. If the solvent is water, the mixture is an aqueous solution. The following solubility rules are for ionic compounds in water as the solvent.

### Solubility Rules for Ionic Compounds in Water

- The **compound is SOLUBLE** if it has:
  1. Li⁺, Na⁺, K⁺, NH₄⁺ ions (ALWAYS!)
  2. Acetate ion (C₂H₃O₂⁻), nitrate ion (NO₃⁻), or perchlorate ion (ClO₄⁻)
  3. Halide ions (X⁻): chloride ion (Cl⁻), bromide ion (Br⁻), or iodide ion (I⁻), but AgX, PbX₂, Hg₂X₂ are **insoluble**
  4. Sulfate ion (SO₄²⁻), but CaSO₄, SrSO₄, BaSO₄ are **insoluble**

- The **compound is INSOLUBLE** if it has:
  5. Carbonate ion, CO₃²⁻, but Rule 1 ions
  6. Chromate ion, CrO₄²⁻, but Rule 1 ions
  7. Phosphate ion, PO₄³⁻, but Rule 1 ions
  8. Sulfide ion, S²⁻, but Rule 1 ions and CaS, SrS, BaS are **soluble**
  9. Hydroxide ion, OH⁻, but Rule 1 ions, Ca(OH)₂, Sr(OH)₂, Ba(OH)₂ are **soluble**

Watch the animations at the following links (sound is needed for the 1st link) to answer the following two questions:


**Question 1:** When sodium chloride dissolves in water explain how the water molecules arrange themselves around the sodium ions. Draw a picture of a sodium ion surrounded by two water molecules.

**Question 2:** Now explain why the water molecules arrange themselves like that around the sodium ions. Describe the intermolecular forces that form between polar water molecules and sodium ions in your answer.
Molecular-Level Representations of Ionic Solutions

For each ionic compound listed below write the chemical formula and draw a representation of the compound in water. Do not show the solvent water molecules. The first two have been completed as examples.

1. potassium sulfate  
   formula: \( K_2SO_4 \)

2. calcium carbonate  
   formula: \( CaCO_3 \)

3. cobalt (III) nitrate  
   formula: \( \text{__________} \)

4. silver carbonate  
   formula: \( \text{__________} \)

5. magnesium sulfate  
   formula: \( \text{__________} \)

6. barium hydroxide  
   formula: \( \text{__________} \)

Chemical Reactions in Solution

Three common types of chemical reactions that take place in solution are:
1. precipitation reactions (McMurry/Fay text, sections 7.3, 7.4)
2. acid base neutralization reactions (McMurry/Fay text, section 7.5)
3. single replacement reactions - redox reactions (McMurry/Fay text, sections 7.6 – 7.8)

To review precipitation reactions watch the following animation:
http://preparatorychemistry.com/precipitation_flash.htm

Question 1: What are the formulas for the four reactant ions involved in the above animation?

Question 2: Why don’t the carbonate and nitrate ions “stick” together in the animation?

Question 3: Write the net ionic equation for the reaction in the animation.
We can write and balance chemicals reactions using molecular equations, ionic equations, and net ionic equations. (McMurry/Fay text, section 7.3) For the following sets of reactants, write and balance the molecular, ionic and net ionic equations (including all phases, using the Solubility Rules on page 1) and draw a molecular-level representation of each reactant and the products in the provided beakers (based on the ionic equation you wrote). Also indicate the type of reaction (No Reaction, Precipitation, or Acid-Base Neutralization). If there is no reaction write NR for the type but you still must complete the reactions and molecular-level drawings. Note that you should draw water molecules if they are part of the reaction, but do not draw solvent water molecules. The first problem (shown below) is already completed as an example.

Example: calcium acetate and ammonium sulfate  Reaction type: __precipitation____

Molecular: \(\text{Ca(CH}_3\text{COO)}_2(aq) + (\text{NH}_4\text{)}_2\text{SO}_4(aq) \rightarrow \text{CaSO}_4(s) + 2 \text{NH}_4\text{CH}_3\text{COO}(aq)\)

Ionic: \(\text{Ca}^{2+}(aq) + 2 \text{CH}_3\text{COO}^-(aq) + 2 \text{NH}_4^+(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{CaSO}_4(s) + 2 \text{CH}_3\text{COO}^-(aq) + 2 \text{NH}_4^+(aq)\)

Net Ionic: \(\text{Ca}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{CaSO}_4(s)\)

1. gold (I) chloride and potassium phosphate  Reaction type: _______________________

Molecular:__________________________________________________________

Ionic:______________________________________________________________

Net Ionic:__________________________________________________________

1. gold (I) chloride and potassium phosphate  Reaction type: _______________________

Molecular:__________________________________________________________

Ionic:______________________________________________________________

Net Ionic:__________________________________________________________
2. perchloric acid and calcium hydroxide

Reaction type: __________________________

Molecular: ________________________________________________________________

Ionic: ____________________________________________________________________

Net Ionic: _________________________________________________________________

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**CHM 152LL: Review of Common Definitions / Calculations**

A brief summary of some useful definitions and relationships is given below; if you need an in-depth review of these concepts, refer to the sections listed below in your *General Chemistry: Atoms First* textbook by McMurray/Fay. The review problems that follow will give you practice on some of the types of calculations that are used in this lab course. In general, when solving numerical problems make sure to show your work including units and express your final answer with the proper number of significant figures.

**Sections 7.6 & 7.7: Redox reactions and oxidation numbers:** A substance is oxidized when it loses one or more electrons. A substance is reduced when it gains one or more electrons. An oxidizing agent causes a substance to be oxidized and is itself reduced. A reducing agent causes a substance to be reduced and is itself oxidized.

\[
\begin{align*}
0 & +1 & +5 & -2 \\
0 & +2 & +5 & -2 \\
\end{align*}
\]

Example: \( \text{Cu (s)} + 2 \text{AgNO}_3 (aq) \rightarrow 2\text{Ag (s)} + \text{Cu(NO}_3)_2 (aq) \)

Oxidation numbers are assigned above each element. To determine what is oxidized or reduced, we must look at how the oxidation numbers changed. Since \( \text{Cu (s)} \) increased from a 0 to +2, it must have lost two electrons. Therefore it was oxidized. \( \text{Ag}^+ \) in \( \text{AgNO}_3 \) decreased from +1 to 0, so it gained one electron and was reduced.

Substances oxidized or reduced are elements. Oxidizing and reducing agents are the full compound used in a reaction (or ion if only a net ionic equation is given).

Oxidized: \( \text{Cu (s)} \)  
Reduced: \( \text{Ag}^+ \text{ in AgNO}_3 (aq) \)

Oxidizing agent: \( \text{AgNO}_3 (aq) \)  
Reducing agent: \( \text{Cu (s)} \)

**Sections 1.9 & 6.3: Molar Mass:** the mass in grams of one mole of a substance; units = g/mol.

Example: Molar mass of KOH = 56.11 g/mol; thus, 56.11 g KOH = 1.00 mole KOH
Section 6.8: Molar Concentration: The concentration of an aqueous solution is often expressed by its molarity (symbolized M), which is defined as the number of moles of solute per liter of solution.

\[
M = \frac{\text{moles solute}}{\text{liters solution}}
\]

Example: For a 0.100 M NaOH solution, there is 0.100 mole of NaOH in 1.00 L of solution.

Conversion factors can be constructed from these and other relations, and they can be “strung together” to solve a problem. For example, to calculate the number of grams of HNO₃ in 25.0 mL of 15.4 M HNO₃ solution, the setup and answer would be:

\[
g \text{HNO}_3 = 25.0 \text{ mL HNO}_3 \text{ soln} \left( \frac{15.4 \text{ moles HNO}_3}{1000 \text{ mL HNO}_3 \text{ soln}} \right) \left( \frac{63.02 \text{ g HNO}_3}{1 \text{ mol HNO}_3} \right) = 24.3 \text{ g HNO}_3
\]

Note: We substituted 1000 mL for 1 L so that mL would cancel out.

Ion Molarities: Soluble ionic compounds completely dissociate into ions in solution; to determine the molarity of individual ions, we must consider the stoichiometric relationship between the ionic compound and the number of ions formed in solution.

Example: What is the molarity of K⁺ ions in a 0.20 M K₂CO₃ solution?

\[
\text{K}_2\text{CO}_3(s) \rightarrow 2\text{K}^+(aq) + \text{CO}_3^{2-}(aq)
\]

\[
\left( \frac{0.20 \text{ mol K}_2\text{CO}_3}{1 \text{ L}} \right) \left( \frac{2 \text{ mol K}^+}{1 \text{ mol K}_2\text{CO}_3} \right) = \frac{0.40 \text{ mol K}^+}{1 \text{ L}} = 0.40 \text{ M K}^+
\]

Section 6.9: Diluting Solutions: Since only water is added during the dilution of aqueous solutions, the moles of solute before dilution (M_{init} \cdot V_{init}) equal the moles of solute after dilution (M_{fin} \cdot V_{fin}) leading to M_{init} \cdot V_{init} = M_{fin} \cdot V_{fin}. To calculate the volume of water added, simply subtract the initial volume from the final volume. Note that any unit of volume can be used in place of liters. Hence, the following formula can be used to solve all dilution problems:

\[
M_1V_1 = M_2V_2
\]

Example: What is the final concentration of 25.0 mL of a 0.15 M NaCl solution that is added to 150.0 mL of water?

\[
M_1 = 0.15 \text{ M}, \quad V_1 = 25.0 \text{ mL}; \quad V_2 = \text{total volume} (25.0 \text{ mL} + 150.0 \text{ mL}) = 175.0 \text{ mL}
\]

\[
M_2 = \frac{M_1V_1}{V_2} = \frac{(0.15 \text{ M} \times 25.0 \text{ mL})}{175.0 \text{ mL}} = 0.021 \text{ M}
\]

Sections 6.11: Titration (Solution Stoichiometry): You must consider the balanced chemical reaction and set up the appropriate conversion factors to solve solution stoichiometry and titration problems.

NOTE: Do not use the dilution formula (M_1V_1 = M_2V_2) for these problems!

Example. How many L of 0.164 M Ca(OH)₂ is needed to neutralize 25.00 mL of 0.458 M HCl solution?

\[
\text{Ca(OH)}_2(aq) + 2 \text{HCl}(aq) \rightarrow \text{CaCl}_2(aq) + 2 \text{H}_2\text{O}(l)
\]

\[
V \text{Ca(OH)}_2 = 25.00 \text{ mL} \left( \frac{1 \text{ L}}{1000 \text{ mL}} \right) \left( \frac{0.458 \text{ moles HCl}}{2 \text{ mol HCl}} \right) \left( \frac{1 \text{ mol Ca(OH)}_2}{1 \text{ L}} \right) \left( \frac{1 \text{ mol Ca(OH)}_2}{0.164 \text{ mol Ca(OH)}_2} \right) = 0.349 \text{ L Ca(OH)}_2
\]
CHM 152LL: Lab Calculation Review Problems

1. Given the balanced equation: \( \text{MnO}_4^{-}(aq) + 5 \text{Fe}^{2+}(aq) + 8 \text{H}^{+}(aq) \rightarrow \text{Mn}^{2+}(aq) + 5 \text{Fe}^{3+}(aq) + 4 \text{H}_2\text{O}(l) \), assign oxidation numbers to each element in the reaction. Then identify what element has been oxidized and which has been reduced. Also identify the oxidizing agent and the reducing agent.

   Oxidized: ____________    Reduced: _______________
   Oxidizing agent: ______________  Reducing agent: ____________

2. a) What is the molarity of strontium ions in a 0.421 M solution of strontium phosphate?

   b) What is the molarity of phosphate ions in this solution?

3. a) How many mL of 5.00 M \( \text{Li}_2\text{CrO}_4 \) are needed to prepare 3.00 liters of 0.250 M \( \text{Li}_2\text{CrO}_4 \)?

   b) How much water is added to prepare this solution?

4. Calculate the molarity of a solution prepared by dissolving 35.8 g of magnesium perchlorate in enough water to yield 275 mL of solution.

5. How many grams of glucose (C\(_6\)H\(_{12}\)O\(_6\)) are required to prepare 1.75 liters of 0.233 M glucose solution?

6. How many grams of AgBr can be formed when 75.0 mL of 0.650 M \( \text{CaBr}_2 \) are reacted with excess \( \text{AgNO}_3 \)? (Write the balanced equation for the reaction.)

7. What is the molarity of a KOH solution if 42.58 mL of KOH are required to neutralize 25.00 mL of 0.1350 M \( \text{H}_3\text{PO}_4 \)? (Write the balanced equation for the reaction.)