Chem E-1a
Friday Review Notes: Chapter 4

Concentration of Solutions:

$$\text{Molarity} = M = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{\text{mol}}{\text{L}}$$

Net-Ionic Equations:

To write a balanced net-ionic equation, follow these three steps:

1. Write the complete balanced equation with phase indications.
2. Write the complete ionic equation by separating into ions the following:
   •All (*aq*) ionic substances
   •All (*aq*) strong acids (i.e. HCl, HBr, HI, HNO_3, $HClO_4$ and H_2SO_4.)
   (Note that H_2SO_4 should be separated into H^+ and HSO_4^-.)
3. Cancel out any ions that appear the same on both sides of the complete ionic equation from step 2 in order to obtain the **net-ionic equation**.

Note: For a simple precipitation reaction, the net ionic reaction can be written simply by knowing the identity of the precipitate: ion (*aq*) + ion (*aq*) → ionic compound (*s*)

Acid-Base Reactions (i.e. Neutralization Reactions):

In order to write a balanced acid/base reaction, remember that the standard template for an acid/base reaction will include the combination of one H^+ with one OH^- to make one H_2O. Thus, to balance an acid/base reaction, you need to have equal numbers of acidic H^+'s and basic OH^-'s. In general:

$H^+ + OH^- \rightarrow H_2O$

We can write these reactions using the common acids that we know. If we don't know the complete chemical formula for an acid, we can abbreviate a monoprotic acid as "HA", a diprotic acid as "H_2A", and a triprotic acid as "H_3A". See the following examples of balanced reactions in order to recognize the standard patterns:

Monoprotic acid + NaOH:

$HCl + NaOH \rightarrow H_2O + NaCl$

$HA + NaOH \rightarrow H_2O + NaA$

Diprotic acid + NaOH:

$H_2SO_4 + 2\ NaOH \rightarrow 2\ H_2O + Na_2SO_4$

$H_2A + 2\ NaOH \rightarrow 2\ H_2O + Na_2A$

Triprotic acid + NaOH:

$H_3PO_4 + 3\ NaOH \rightarrow 3\ H_2O + Na_3PO_4$

$H_3A + 3\ NaOH \rightarrow 3\ H_2O + Na_3A$

Try writing out a similar set of reactions, but using the base $Ba(OH)_2$:

$2\ HCl + Ba(OH)_2 \rightarrow 2\ H_2O + BaCl_2$

$H_2SO_4 + Ba(OH)_2 \rightarrow 2\ H_2O + BaSO_4$

$2\ H_3PO_4 + 3\ Ba(OH)_2 \rightarrow 6\ H_2O + Ba_3(PO_4)_2$

Titrations:

A titration is a method of "counting" the **quantity** of a particular substance in a sample by reacting the substance with a solution of known concentration. In essence, a titration is a laboratory technique by which we can "count" the number of moles of something in a sample.

The "endpoint" or "equivalence point" of a titration is the point at which **stoichiometrically equivalent** quantities of the reactants have been added.

By knowing the number of moles of one reactant that were added, you can use the stoichiometric coefficients in the balanced chemical reaction to determine the number of moles of the other reactant that were initially present. If the titration involved an acid/base reaction, the "endpoint" or "equivalence point" could also be referred to as the point at which the original acid or base had been "completely neutralized".

Oxidation Numbers:

The following are the rules from the lecture notes for assigning oxidation numbers. Follow these rules **in the order listed**.

1. Pure elements: For any atom in elemental form the oxidation number is zero.

2. Monatomic ions: For monatomic ions in ionic compounds the oxidation number equals the charge on the ion. Group I and Group II ions are always +1 and +2, respectively. Oxidation numbers of transition metal monatomic ions can vary, but can often be determined by recognizing the identity and charge of the anion they are bonded to. Monatomic anions will have oxidation numbers equal to the charge on the ion.

3. Fluorine: always –1 in a compound.

4. Hydrogen: +1 in a compound, unless violated by rule 2 or rule 3 above.

5. Oxygen: –2 in a compound, unless violated by rules 2, 3, or 4 above.

6. For neutral species the oxidation numbers add to zero, for ions they add to the charge on the ion.

Using RICE Tables:

RICE tables provide a simple method of "book-keeping" to track the stoichiometric changes in moles of reactants and products in a chemical reaction. In essence, you could refer to a RICE table as "stoichiometry by spreadsheet"! At this point in the course, using RICE tables is not absolutely necessary, but you may find it helpful. It can be particularly helpful if you are asked to calculate the quantities of several different products (and reactants) that were produced (or remain unreacted) after a chemical reaction is complete. As the course progresses (after the first exam) RICE tables will become necessary to solve some of the stoichiometry problems we will see.

It is difficult to explain the process of setting up a RICE table, and some people have slightly different approaches. Listed below is one general technique for setting up and using a RICE table:

1. Write the balanced chemical reaction, and set up the RICE table with the rows labeled "R" for Reaction, "I" for Initial, "C" for Change, and "E" for End. Decide the units you will use for the RICE table. At this point in the course, we will only do RICE tables using "number of moles" as our units.

2. Fill in the "I" row with the initial numbers of moles of each reactant and product as indicated in the problem. If something (like a product) is not mentioned as being present, then there is none of it there initially. Use only information that is given in the problem, and ignore the coefficients in the balanced balanced chemical reaction.

3. Fill in the "C" row based only on the coefficients in the balanced reaction. Do this by writing a "–x" or "+x" for each reactant and product, with the x preceded in each case by the stoichiometric coefficient from the balanced chemical reaction. When doing this, completely ignore the numbers in the "I" row or any other information given in the problem.

4. Solve for x and plug in the value of x. If the reaction goes to completion with a limiting reagent (i.e. your are told the it "reacts completely" or "goes to completion" or "reacts to completion"), then solve for x by dividing each number in the "I" row for each reactant by the stoichiometric coefficient in front of that reactant. The smallest number you get is the value of x.
(If the reaction doesn't go to completion, then you would need to use other information given in the problem to solve for x. We will see many examples of this type of problem later in the course.)

5. Add down each column to get the "E" row. This will give you the number of moles of every reactant and product that is present after the reaction is complete.

Using Dimensional Analysis With Stoichiometry