**Chemistry 1314 – Lecture Outline and Notes**
Text: Chemistry – The Central Science, 10$^{th}$ edition, Brown, LeMay and Bursten

Chapter 4. Aqueous Reactions and Solution Stoichiometry

Solutions in which water is the dissolving medium are called aqueous solutions.

**General Properties of Aqueous Solutions**
If we have a solution (a homogeneous mixture), the substance present in greater quantity is the **solvent**. The substance that *dissolves* in it is called the **solute**. For example, NaCl dissolved in water, the water is the solvent and the sodium chloride the solute.

Electrolytic properties
**Electrolyte** – solutions containing ions (such as NaCl)
**Nonelectrolyte** – a substance that does not form ions (such as sucrose ($C_{12}H_{22}O_{12}$))

Ionic Compounds in Water
Ionic compounds *dissociate* into their individual ions as it dissolves. $NaCl \rightarrow Na^+ + Cl^-$
The water molecules surround the ions as NaCl dissolves.
Given a substance, we can usually predict its ions – such as $Na^+$ and $Cl^-$ from NaCl

Molecular Compounds in Water
The molecules usually stay intact and are dispersed throughout the solution. Such as sucrose is dispersed throughout sugar water.
A special case arises with the acids, such as HCl – it dissolves into $H^+$ and $Cl^-$

Strong and weak electrolytes
A strong electrolyte completely dissociates into its ions

$$HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$$

A weak electrolyte does not completely dissociates

$$HC_2H_3O_2\ (aq) \leftrightarrow H^+\ (aq) + C_2H_3O_2^-\ (aq)$$

Notice that this reaction has a double arrow, this says that some react and goes to product, but we can also say that some of the products recombine to form the reactant. A point is established where they are in **chemical equilibrium**.

**Precipitation Reactions**
Reactions that form an insoluble product are called **precipitation reactions**.

$$Pb(NO_3)_2\ (aq) + 2KI\ (aq) \rightleftharpoons PbI_2\ (s) + 2KNO_3\ (aq)$$

Certain pairs of oppositely charged ions attract so strongly that they form insoluble ionic solids.

How do we know which compounds will be insoluble. Again, some memorization might need to be done.
Become familiar with **Table 4.1 p. 127**.

Are the following soluble or insoluble?

1. cobalt(II) hydroxide [$Co(OH)_2$]
2. barium nitrate [$Ba(NO_3)_2$]
3. ammonium phosphate [$(NH_4)_3PO_4$]

To determine what will be soluble or insoluble, we have to be familiar with what ions will be present in the solution and how they can interact.
Example:

$$Mg(NO_3)_2\ (aq) + 2NaOH\ (aq) \rightarrow Mg(OH)_2\ (s) + 2NaNO_3\ (aq)$$

We see that the ions in the solution will be $Mg^{2+}$, $NO_3^-$, $Na^+$, and $OH^-$. If we examine the possible combinations [$Mg(NO_3)_2$, $NaOH$, $Mg(OH)_2$, and $NaNO_3$], we see that $Mg(OH)_2$ would be insoluble and precipitate out.

The above reaction has the form of

$$AX + BY \rightarrow AY + BX$$

The ionic partners are switched. This is called an **exchange reaction**, or **metathesis reaction**

Example:
What compound precipitates when solutions of $Fe_2(SO_4)_3$ and LiOH are mixed?

Write a balanced equation for this reaction:

Will a precipitate form when solutions of $Ba(NO_3)_2$ and KOH are mixed?

**Ionic equations**

I give you the following balanced equation:

$$Pb(NO_3)_2\ (aq) + 2KI\ (aq) \rightarrow PbI_2\ (s) + 2KNO_3\ (aq)$$

This is an example of a **molecular equation** because it shows the complete chemical formula of each component.

But, this reaction is taking place in water, so the *soluble* components are in their ionic form. So we have:

$$Pb^{2+}\ (aq) + 2NO_3^-\ (aq) + 2K^+\ (aq) + 2I^-\ (aq) \rightarrow PbI_2\ (s) + 2K^+\ (aq) + 2NO_3^-\ (aq)$$

This is known as the **complete ionic equation**

We see that the $NO_3^-$ and $K^+$ are in the same state on both sides of the equation. We can eliminate those and end up with the following:

$$Pb^{2+}\ (aq) + 2I^-\ (aq) \rightarrow PbI_2\ (s)$$

This equation is the **net ionic equation**, and it shows us only the components that are doing something. The $NO_3^-$ and $K^+$ are just standing around doing nothing for the reaction – they are **spectator ions**.

Example:
What is the net ionic equation for the precipitation reaction that occurs on mixing solutions of silver nitrate and potassium phosphate?

The net ionic equation shows us only the components that are involved in the reaction.

**Acid-Base Reactions**
**Acids** are substances that form $H^+$ ions in aqueous solutions. $H^+$ only contains a proton, thus acids are called *proton donors***.**

Acids which yield only 1 $H^+$ ion, such as HCl (hydrochloric acid) and $HNO_3$ (nitric acid), are called monoprotic (one proton).
Acids that yield 2 $H^+$ ions, such as $H_2SO_4$ (sulfuric acid), are diprotic. $H_2SO_4$ releases its 2 $H^+$ protons in 2 steps:

$$H_2SO_4\ (aq) \rightarrow H^+\ (aq) + HSO_4^-\ (aq)$$
$$HSO_4^-\ (aq) \rightleftharpoons H^+\ (aq) + SO_4^{2-}\ (aq)$$

Note that $H_2SO_4$ is a strong electrolyte (and, thus, a strong acid) but only the first ionization is complete, the second will reach a chemical equilibrium so that the final solution will contain $H^+$ (aq), $HSO_4^-$ (aq), and $SO_4^{2-}$ (aq).

**Bases** accept a proton ($H^+$). Bases produce hydroxide ions ($OH^-$) when they dissolve in water. Examples are NaOH, KOH, $Ca(OH)_2$.
Another example is $NH_3$, which is a weak base. When it enters water, it forms an $OH^-$ ion:

$$NH_3\ (aq) + H_2O\ (l) \rightleftharpoons NH_4^+\ (aq) + OH^-\ (aq)$$

[NOTE: this reaction is reversible, only about 1% of $NH_3$ forms $NH_4^+$ and $OH^-$, therefore $NH_3$ is a weak electrolyte and, thus, a weak base.]

Acids and bases that are strong electrolytes (completely ionized in solution) are **strong acids** and **strong bases** respectively. The stronger the acid (or base), the more reactive.

STRONG ACIDS:
Hydrogen with a halogen (group 7A) except F, are strong: HCl (hydrochloric), HBr (hydrobromic), HI (hydroiodic)
Some very common acids are also there: $HClO_3$ (chloric), $HClO_4$ (perchloric), $HNO_3$ (nitric), $H_2SO_4$ (sulfuric).

STRONG BASES:
Group 1A metal hydroxides: LiOH, NaOH, KOH, RbOH, CsOH
Heavy group 2A metal hydroxides: $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$

**Strong and weak electrolytes [Table 4.3, p. 133]**
Strong electrolytes: all ionic compounds and the strong acids
Weak electrolytes: weak acids and weak bases ($NH_3$)
Nonelectrolytes: all others (such as glucose ($C_6H_{12}O_6$), methanol $CH_3OH$)

**Neutralization Reactions and Salt**
When acidic and basic solutions are mixed, they are neutralized – a **neutralization reaction**. These reactions make water and a salt.
Example:

$$HCl\ (aq)\ +\ NaOH\ (aq)\ \rightarrow\ H_2O\ (l)\ +\ NaCl\ (aq)$$

This is a matathesis (or exchange) reaction, but the net ionic equation is:

$$H^+\ (aq)\ +\ OH^-\ (aq)\ \rightarrow\ H_2O\ (l)$$

This summarizes the net reaction between a strong acid and strong base.

Example:
(a) Write a balanced equation for the reaction of carbonic acid ($H_2CO_3$) and potassium hydroxide (KOH). (b) Write the next ionic equation for this reaction.
Answer:

Sometimes, an acid and base will react and form a gas. We have all either made coke bottle rockets or model volcanoes and used baking soda ($NaHCO_3$) and vinegar ($HC_2H_3O_2$).
$NaHCO_3$ is commonly called sodium bicarbonate, but officially sodium hydrogen carbonate
$HC_2H_3O_2$ is acetic acid, a weak acid

$$HC_2H_3O_2\ (aq) + NaHCO_3\ (aq) \rightarrow H_2CO_3\ (aq) + NaC_2H_3O_2\ (aq)$$

HOWEVER, $H_2CO_3$ will decompose:

$$H_2CO_3\ (aq) \rightarrow H_2O\ (l) + CO_2\ (g)$$

We can express the overall reaction as:

$$HC_2H_3O_2\ (aq) + NaHCO_3\ (aq) \rightarrow NaC_2H_3O_2\ (aq) + H_2O\ (l) + CO_2\ (g)$$

With a net ionic equation of:

$$HC_2H_3O_2\ (aq) + HCO_3^-\ (aq) \rightarrow C_2H_3O_2^-\ (aq) + H_2O\ (l) + CO_2\ (g)$$

The $CO_2$ is what causes the foaming, and since $CO_2$ is a gas, a small amount will fill the coke bottle quickly with strong pressure that will have to be released for the launch.

**Oxidation-Reduction Reactions:**
Reactions in which electrons are transferred between reactants are called **oxidation-reduction**, or **redox**, **reactions**.

**Oxidation** – loss of electrons by a substance
**Reductions** – gain of electrons by a substance

The word reduction can be confusing, because we are gaining electrons. We can think of it as "reducing" the charge (getting more negative with more negative electrons).

The name oxidation comes from one particular reaction – metals with oxygen in the air and corrodes. The metal is oxidized in this reaction, it looses electrons.

$$2Ca\ (s) + O_2\ (g) \rightarrow 2CaO\ (s)$$

The Ca goes to $Ca^{2+}$ (looses 2 electrons, so is oxidized)
The $O_2$ goes to $O^{2-}$ (gains electrons and is reduced)

But, the reaction does not have to involve oxygen:

$$Ca\ (s) + 2H^{+}\ (aq) \rightarrow Ca^{2+}\ (aq) + H_2\ (g)$$

**Oxidation Numbers** – the charge on an atom if it is a monoatomic ion; otherwise, it is assigned by a set of *rules*:

1. Atom in *elemental form* = oxidation number is zero
2. For a *monoatomic ion* = charge on the ion
3. *Nonmetals*:
   a) *oxygen* is usually -2 (major exception is in peroxides, $O_2^{2-}$, where it is -1)
   b) *hydrogen* is +1 when bound to nonmetals, and -1 when bound to metals
   c) *halogens* (group 7A)
      1) Fluorine is always -1
      2) The others are -1 for most cases, except when bound with oxygen where they have a positive number
4. The *sum of the oxidation numbers* equals zero for a neutral compound. It is equal to the charge if it is part of an ion.

Examples:
What are the oxidation numbers for the following boldfaced elements?

A) $\mathbf{P}_2O_5$

B) Na**H**

C) $\mathbf{Cr}_2O_7^{2-}$

D) **Sn**$Br_4$

E) Ba$\mathbf{O}_2$

There are many kinds of redox reactions, such as a combustion reaction. We are going to consider redox reactions between metals and either acids or salts. The reactions are **displacement reactions** of the form:

$$A + BX \rightarrow AX + B$$

Examples:

$Zn (s) + 2HBr (aq) \rightarrow ZnBr_2 (aq) + H_2 (g)$

Let's look at the oxidation number of the above reactions:

$Zn (s) + 2HBr (aq) \rightarrow ZnBr_2 (aq) + H_2 (g)$

Ox. # (0) (+1) (-1) (+2) (-1) (0)

We can say that the Zn is oxidized, going from oxidation number 0 to +2, a loss of electrons.

The H has gained electrons, one for each hydrogen, or 2 for $H_2$ – it is reduced

$Mn (s) + Pb(NO_3)_2 \rightarrow Mn(NO_3)_2 (aq) + Pb (s)$

Ox. # (0) (+2) (-1) (+2) (-1) (0)

Again, Mn looses 2 electrons, it is oxidized

Pb gains 2 electrons and is reduced.

Example:

a) Write the balanced molecular and net ionic equations for the reaction between magnesium and cobalt(II) sulfate. b) What is oxidized and what is reduced in the reaction?

Answer:

How do we know if in the above reactions that the Zn will be oxidized by $H^+$, or Mn by $Pb^{2+}$, or Mg by $Co^{2+}$? There is an activity series for the metals (**TABLE 4.5, p. 143**)

The lower in the series, the less likely it will be oxidized. If a metal is above another, it will be oxidized by that metal. We see that Zn is above H, so $H^+$ will oxidize Zn. The same is true for Mn is above Pb, and Mg is above Co.

Example:

Which of the following metals will be oxidized by $Pb(NO_3)_2$: Zn, Cu, Fe?

Answer:

**Concentrations of Solutions**
Most of the reactions that you will experience are conducted in an aqueous environment. We need a way of saying how much material we have in a solution, or its **concentration**. The more solute that is dissolved, the more concentrated the resultant solution.

**Molarity**:
We express the concentration of a solution as **molarity** (**M**). We say the solution is X **molar**.
Molarity = **moles** solute / volume of solution in **liters**

For example, a 1.00 M solution contains 1.00 mol of solute in 1.00 L of solution.

Example:
Calculate the molarity of a solution made by dissolving 5.00 g of glucose ($C_6H_{12}O_6$) in sufficient water to form exactly 100 ml of solution.
Answer:

What if we have an ionic solution of a salt?
Example:
What is the concentration of $K^+$ ions in a 0.015 M solution of potassium carbonate?
Answer:

We can interconvert between molarity, moles, and volume.
Example:
a) How many grams of $Na_2SO_4$ are there in 15 ml of 0.50 M $Na_2SO_4$? b) How many milliliters of 0.50 M $Na_2SO_4$ solution are needed to provide 0.038 mol of this salt?
Answer:

In the lab, we often make a concentrated stock solution, and we will take a sample out and dilute to the amount needed. Suppose we have a 1.00 M solution of NaCl and need to make a 250 ml of a 0.100 M solution. First, we can determine how many moles of NaCl is in the 0.100 M solution.
0.250 L (0.100 mol/ 1 L) = 0.0250 mol NaCl
Now, we can determine what volume of our stock solution contains 0.0250 mol NaCl
0.0250 mol NaCl (1 L/1 mol) = 0.025 L = 25.0 ml NaCl

If we remove 25.0 ml of NaCl, which contains 0.0250 mol NaCl, and put it into a flask and then dilute it to 250 ml, we will obtain our 0.100 M NaCl solution. Note that when we dilute the 25.0 ml sample, we are changing the volume, not the amount of NaCl, so our final 250 ml sample still has 0.0250 mol of NaCl.

We could have done this by realizing the following relationship, as the moles will be constant:

$$M_{conc} \times V_{conc} = M_{dil} \times V_{dil}$$

And since this is a relationship, we can use L or ml or M or mM, or any other combination <u>as long as the same unit is used on each side!!!</u>

So, from above:

$$(1.00\ M) \times V_{conc} = (0.100\ M) \times (250\ ml)$$

$$V_{conc} = 25.0\ ml$$

Examples:
a) What volume of 2.50 M lead nitrate solution contains 0.0500 mol of $Pb^{2+}$? b) How many milliliters of 5.0 M $K_2Cr_2O_7$ solution must be diluted to prepare 250 ml of 0.10 M solution? c) If 10.0 ml of a 10.0 M stock solution of NaOH is diluted to 250 ml, what is the concentration of the resulting solution?
Answers:

**Titrations**
A titration is when a known volume of an unknown solution has a known solution, a **standard solution**, added to it and the amount is noted. The end point, or the **equivalence point**, is indicated by some type of **indicator**. We might be familiar with litmus paper that turns red in the presence of acids and blue in the presence of bases.
So, with a titration we have volume of the unknown but not the concentration. The standard solution we have the concentration and we measure the volume to reach the equivalence point. We can use this information to determine the concentration of the unknown.

Example:
A sample of an iron ore is dissolved in acid, and the iron is converted to $Fe^{2+}$. The sample is then titrated with 47.20 ml of 0.02242 M $MnO_4^-$ solution. The oxidation-reduction reaction that occurs during the titration is as follows: $MnO_4^-\ (aq) + 5Fe^{2+}\ (aq) + 8H^+\ (aq) \rightarrow Mn^{2+}\ (aq) + 5Fe^{3+}\ (aq) + 4H_2)\ (l)$. a) How many moles of $MnO_4^-$ were added to the solution? b) How many moles of $Fe^{2+}$ were in the sample? c) How many grams of iron were in the sample? d) If the sample had a mass of 0.8890 g, what is the percentage of iron in the sample?
Answer: