**Reactions in Aqueous Solutions: Metathesis Reactions and Net Ionic Equations**

To become familiar with writing equations for metathesis reactions, including net ionic equations.

**Materials**

**Part A:** well plate and dropper bottles with necessary solutions

**Part B:** evaporating dish, thermometer, Bunsen burner/hose, magnifier, ring stand and ring,

wire gauze, two 100mL beakers, one 600mL beaker, filter paper, funnel

In molecular equations for many aqueous reactions, cations and anions appear to exchange partners. These reactions conform to the following general equation: AX + BY → AY + BX [1]

These reactions are known as metathesis reactions. For a metathesis reaction to lead to a net change in solution, ions must be removed from the solution. In general, three chemical processes can lead to the removal of ions from solution, thus serving as a *driving force* for metathesis to occur:

1. The formation of a precipitate

2. The formation of a weak electrolyte or nonelectrolyte

3. The formation of a gas that escapes from solution

**Formation of a Precipitate**

The reaction of barium chloride with silver nitrate is a typical example:

BaCl2(aq) + 2AgNO3(aq) → Ba(NO3)2(aq) + 2AgCl(s) [2]

This form of the equation for this reaction is referred to as the *molecular equations*. Since we know that the salts BaCl2, AgNO3, Ba(NO3)2 are strong electrolytes and are completely dissociated in solution, we can more realistically write the equation as follows:

Ba2+(aq) + 2Cl−(aq) + 2Ag+(aq) + 2NO3−(aq) → Ba2+(aq) + 2NO3−(aq) + 2AgCl(s) [3]

This form, in which all ions are shown, is known as the complete ionic equation. Reaction [2] occurs because the insoluble substance AgCl precipitates out of solution. The other product, barium nitrate, is soluble in water and remains in solution. We see that Ba2+ and NO3− ions appear on both sides of the equation and thus do not enter into the reaction. Such ions are called *spectator* *ions*. If we eliminate or omit them from both sides (and reduce the coefficients), we obtain the net ionic equation:

Ag+(aq) + Cl−(aq) → AgCl(s) [4]

This equation focuses our attention on the salient feature of the reaction: the formation of the precipitate AgCl. It tells us that solutions of any soluble Ag+ salt and any soluble Cl−salt, when mixed, will form insoluble AgCl. When writing net ionic equations, remember that only strong electrolytes are written in the ionic form. Solids, gases, nonelectrolytes, and weak electrolytes are written in the molecular form. Frequently the symbol (aq) is omitted from ionic equations. The symbols (g) for gas and (s) for solid should not be omitted. Thus, Equation 4 can be written as

Ag+ + Cl− → AgCl(s) [5]

Consider mixing solutions of KCl and NaNO3. The ionic equation for the reaction is

K+(aq) + Cl−(aq) + Na+(aq) + NO3−(aq) → K+(aq) + Cl−(aq) + Na+(aq) + NO3−(aq) [6]

Because all the compounds are water-soluble and are strong electrolytes, they have been written in the ionic form. They completely dissolve in water. If we eliminate spectator ions from the equation, nothing remains. Hence, there is no reaction:

K+(aq) + Cl−(aq) + Na+(aq) + NO3−(aq) → no reaction [7]

Metathesis reactions occur when a precipitate, a gas, a weak electrolyte, or a nonelectrolyte is formed. The following equations are further illustrations of such processes.

**Formation of a Gas**

Molecular equation: 2HCl(aq) + Na2S(aq) → 2NaCl(aq) + H2S(g)

Complete ionic equation:  2H+(aq) + 2Cl−(aq) + 2Na+(aq) + S2−(aq) → 2Na+(aq) + 2Cl−(aq) + H2S(g)

Net ionic equation:  2H+(aq) + S2−(aq) → H2S(g) or 2H+ + S2− → H2S(g)

**Formation of a Weak or Non Electrolyte**

Molecular equation: NaF(aq) + HBr(aq) ⇄ HF(aq) + NaBr(aq)

Complete ionic equation: Na+(aq) + Br−(aq) + H+(aq) + Br−(aq) ⇄ HF(aq) + Na+(aq) + Br−(aq)

Net ionic equation: H+(aq) + F−(aq) ⇄ HF(aq) (HF is a weak-electrolyte)

Molecular equation: HNO3(aq) + NaOH(aq) → H2O(l) + NaNO3(aq)

Complete ionic equation: H+(aq) + NO3−(aq) + Na+(aq) + OH−(aq) → H2O(l) + Na+(aq) + NO3−(aq)

Net ionic equation: H+(aq) + OH−(aq) → H2O(l) (water is a non-electrolyte)

In order to decide if a reaction occurs, we need to be able to determine whether or not a precipitate, a gas, a nonelectrolyte, or a weak electrolyte will be formed. The following brief discussion is intended to aid you in this regard. The solubility rules are at the end of the discussion should be consulted while performing this experiment.

The common gases formed during metathesis reactions are CO2, SO2, H2S, and NH3. Carbon dioxide and sulfur dioxide may be regarded as resulting from the *decomposition* of their corresponding *weak acids*, which are initially formed when carbonate and sulfite salts are treated with acid:

K2CO3(aq) + 2HI(aq) → H2CO3(aq) + 2KI(aq) (the carbonic acid decomposes) H2CO3(aq) → H2O(l) + CO2(g)

the molecular reaction should be written as: K2CO3(aq) + 2HI(aq) → 2KI(aq) + H2O(l) + CO2(g)

similarly sulfurous acid decomposes as: H2SO3(aq) → H2O(l) + SO2(g)

Ammonium salts form ammonia gas when they are treated with strong bases:

NH4+(aq) + OH−(aq) ⇄ NH3(aq) + H2O(l)

Which are the weak electrolytes? The easiest way of answering this question is to *identify* all of the ***strong*** electrolytes, and if the substance *does not* fall in that category then it is a weak electrolyte. Note, water is a nonelectrolyte. Strong electrolytes are summarized at the end of this section.

In the first part of this experiment, you will study some metathesis reactions. In some instances it will be very evident that a reaction has occurred, whereas in others it will not be so apparent. In the doubtful case, use the guidelines above to decide whether or not a reaction has taken place. You will be given the names of the compounds to use but not their formulas. This is being done deliberately to give practice in writing formulas from names.

In the second part of this experiment, you will study the effect of temperature on solubility. The effect that temperature has on solubility varies from salt to salt. We concluded that mixing solutions of KCl and NaNO3 resulted in no reaction (see Equations 6 and 7). What would happen if we cooled such a mixture? The solution would eventually become saturated with respect to one of the salts, and crystals of that salt would begin to appear as its solubility was exceeded. Examination of Equation 6 reveals that crystals of any of the following salts could appear initially: KNO3, KCl, NaNO3, or NaCl. Consequently, if a solution containing Na+, K+, Cl−, and NO3− ions is evaporated at a given temperature, the solution becomes more and more concentrated and will eventually become saturated with respect to one of the four compounds. If evaporation is continued, that compound will crystallize out, removing its ions from solution. The other ions will remain in solution and increase in concentration. A graph of the solubilities of the four salts given at the end of the lab.

***General Solubility rules that apply to water solutions:***

(1) All **alkali metal** (Group IA) and **ammonium** compounds are **soluble**.  
(2) All **acetate, perchlorate, chlorate,** and **nitrate** compounds are **soluble**.  
(3) **Silver, lead,** and **mercury(I)Hg22+** compounds are **insoluble**.  
(4) **Chlorides, bromides,** and **iodides** are **soluble**. **Fluorides** except 2A metals & Pb+2.  
(5) **Carbonates, oxides, phosphates, silicates,** and **sulfides** are **insoluble**.

(6) **Hydroxides** are **insoluble** except **calcium, barium, strontium**.

(7) **Sulfates** are **soluble** except for **calcium, barium, strontium**.

These rules are applied in the order given. For example, PbSO4 is insoluble because rule 3 comes before rule 7. In like manner, AgCl is insoluble because rule 3 takes precedence over rule 4. NaOH is soluble due to rule 1 over 6.

**Strong Electrolytes:**

all soluble salts; strong acids - HCl, HBr, HI, H2SO4, HNO3, HClO4, HClO3

strong bases *(soluble)* – group 1A metal hydroxides and heavy group IIA hydroxides (Ca, Sr, Ba)

**Weak Electrolytes:**

Weak acids; weak bases; slightly soluble salts

**Experimental Procedure**

**Part A: Metathesis Reactions**

1. The report sheet lists 16 pairs of chemicals that are to be mixed. Use about 5 drops of the reagents to be combined as indicated on the report sheet. (add a few more drops it the reaction is not apparent)
2. Mix the solutions in a well plate and record your observations on the report sheet. If there is no reaction, write N.R. (The reactions need not be carried out in the order listed. In order to reduce congestion at the supply table start with chemicals that are available. Return the dropper bottles as you use them. Rinse and clean the well plate in your sink when you are finished.
3. *Note: pairs 13-16 should be mixed in small test tubes since the tubes may need to be heated to observe any odors (gases) produced*. (#13 does NOT need to be heated)

**Part B: Solubility, Temperature and Crystallization**

1. Place about 8.5 g of sodium nitrate and about 7.5 g of potassium chloride in a 100-mL beaker *(zero the beaker each time)* and add 25 mL of water. Warm the mixture, stirring, until the solids completely dissolve.
2. Assuming a volume of 25mL for the solution, calculate the molarity of the solution with respect to NaNO3, KCl, NaCl, and KNO3, and record these molarities on your report form. ***(1)***
3. Cool the solution to about 10°C by placing the beaker in ice water in a 600-mL beaker and stir the solution carefully with a thermometer, being careful not to break it. When no more crystals form, at approximately 10°C, filter the cold solution quickly and allow the filtrate to drain thoroughly into an evaporating dish. Dry the crystals between two dry pieces paper towel.
4. Examine the crystals with a magnifier. Describe the shape of the crystals—that is, needles, cubes, plates, rhombs, and so on your report form. ***(2)***
5. Based upon the solubility graph, which compound crystallized out of solution and write that in the appropriate place on your report form. ***(3)***
6. Evaporate the filtrate to about half of its volume using a Bunsen burner and ring stand. A second crop of crystals should form. Record the temperature and rapidly filter the hot solution, collecting the filtrate in a clean 100-mL beaker. ***(4)***
7. Dry the second batch of crystals between two pieces of paper towel and examine their shape. ***(5)*** Compare their shape with the first batch of crystals. Based upon your solubility graph, what is this substance? ***(6)***
8. Finally, cool the filtrate to 10°C while stirring carefully with a thermometer to obtain a third crop of crystals. Observe their shapes and compare them with those of the first and second batches. ***(7)***

What compound is the third batch of crystals? ***(8)***  Clean and rinse the beakers, evaporating dish and thermometer. Allow the ring stand to cool and return the equipment.