Experiment Twelve

Ionic reactions

Introduction:

Double replacement reactions (ionic reactions) can be subdivided into three groups:

1. Precipitation reactions
2. Neutralization reactions
3. Unstable product (gas evolution) reactions

These reactions are characterized by the type of products they produce. For example, precipitation reactions produce an insoluble salt called a precipitate. Formation of precipitates can be predicted using solubility rules given on the next page. Neutralization reactions always produce water as a product, usually along with a soluble salt. Gas evolution reactions produce a gas due to the unstability of one of the products formed and its subsequent decomposition into a gas.

The most common method to represent a chemical reaction is a ***balanced chemical equation***. This equation is commonly referred to as an ***overall*** or ***molecular*** equation. This equation describes the chemical reaction in general terms, but it may not describe the precise chemical changes that occur at the molecular or ionic levels. This is because the specie that actually experiences the chemical change may be an ion, which is only part of the compound from which it is obtained.

In order to distinguish between substances that are present as ions in solution and those that exist as molecules or as ionic solids, an ***ionic equation*** is written. A ***complete ionic equation*** includes all chemical species actually present in a reaction except the solvent water, which is neither a reactant nor a product, but merely a solvent. Not all species in the reaction undergo chemical change. Substances which are present but experience no chemical change are called **spectator ions.**

A ***net ionic equation*** is an equation in which spectator ions are removed from the total ionic equation. The net ionic equation shows only those reactants that are actually consumed and those products that are actually formed in the reaction.

Solubility Rules

Rule 1. All compounds containing group I A cations (Alkali metals) or the ammonium ion (NH_4^+) are soluble.

Rule 2. All nitrate, acetate and chlorate compounds are soluble.

Rule 3. All sulfate salts are soluble, except the sulfate salts of Pb, Ag, Hg, Ba, Sr and Ca.

Rule 4. All chlorides, bromides and iodides are soluble except for the compounds of these salts that contain Ag, Pb and Hg.

Rule 5. All carbonates, sulfites and phosphates are insoluble. The exceptions to this rule are the compounds that contain NH_4^+, and Alkali metal cations (Group I A). See Rule 1.

Rule 6. All hydroxides are insoluble except those formed with NH_4^+, Group I A metals and Barium. See rule 1.

Rule 7. All dichromate salts are insoluble except the salts that contain Alkali metals and NH_4^+. See rule 1.

<u>Procedure:</u>

For each reaction described in parts 1 and 2, follow the steps listed below

- Mix 10 drops of each of the 2 solutions indicated in a test tube.
- Agitate the two reagents by gently tapping the side of the tube and let the mixture sit for one to two minutes. Observe the test tube for any signs of formation of a solid (precipitate).
- If a precipitate appears, use the solubility rules to determine the identity of the precipitate (ppt).
- Write a balanced molecular equation, using state designations. Identify the precipitate and its color and record your observations.
- In addition to the balanced equation, where requested, write the complete and net ionic equations.
- If no reaction occurs, write "NO REACTION."

Example:

Mix and briefly agitate 10 drops of silver nitrate and potassium dichromate

Color of ppt: Ruby red Formula of ppt: $Ag_2Cr_2O_7$

Molecular equation: $2\ AgNO_3\ (aq) + K_2Cr_2O_7\ (aq) \rightarrow Ag_2Cr_2O_7\ (s) + 2\ KNO_3\ (aq)$

Complete ionic equation: $2\ Ag^+ + \cancel{2\ NO_3} + \cancel{2\ K^+} + Cr_2O_7^{2-} \rightarrow Ag_2Cr_2O_7\ (s) + \cancel{2\ K^+}\ \cancel{2\ NO_3}^-$

Net ionic equation: $2\ Ag^+ + Cr_2O_7^{2-} \rightarrow Ag_2Cr_2O_7\ (s)$

For each reaction described in parts 3and 4, follow the steps listed below

- Mix 10 drops of each of the 2 solutions indicated in a test tube.
- Agitate the two reagents by gently tapping the side of the tube and let the mixture sit for one to two minutes. Observe the test tube for any signs of occurrence of the reaction.
- If a reaction occurs, write balanced molecular equation, with state designations.

Part 1

Mix and briefly agitate 10 drops of each reagent

a) Lead (II) nitrate and potassium chromate

b) Sodium chloride and silver nitrate

c) Lead (II) nitrate and potassium iodide

d) Barium chloride and potassium chromate

Part 2

Mix and briefly agitate 10 drops of each reagent

a) Sodium sulfate and barium chloride

b) Copper (II) sulfate and barium nitrate

c) Aluminum sulfate and barium nitrate

Part 3

Mix and briefly agitate 10 drops of each reagent

a) Sodium hydroxide and hydrochloric acid

b) Ammonium hydroxide and sulfuric acid

c) Sodium hydroxide and nitric acid

Part 4

Add 10 drops of hydrochloric acid to a small amount following solid reagents

a) Sodium carbonate

b) Sodium sulfite