Types of Chemical Reactions

Objectives
- Perform and observe the results of a variety of chemical reactions.
- Become familiar with the observable signs of chemical reactions.
- Identify the products formed in chemical reactions and predict when a reaction will occur.
- Write balanced equations for the reactions studied.
- Use the results from the single replacement reactions to devise a partial activity series.

Background
Matter undergoes three kinds of change: physical, chemical, and nuclear. The composition of a chemical substance is not altered by physical changes such as freezing and evaporation. Chemical changes, or reactions, result in the formation of new substances when bonds are formed and/or broken. Some relatively simple but common types of chemical reactions are illustrated in this experiment. Examples and descriptions of each type appear below.

There are often accompanying observable changes as a chemical reaction occurs. These include:
- Color change.
- Formation of a precipitate—noted as the formation of a cloudy solution, formation of a gel, or an obvious solid.
- Evolution of a gas—noted as bubbling in the solution.
- Appearance or disappearance of distinct separation between two or more liquids.
- Evolution of heat—noted as a temperature increase.
- Absorption of heat—noted as a temperature decrease.
- Plating out of one metal on another.
- Decomposition, pitting, or the disappearance of a solid metal

One or more of these changes may occur in the reactions that are performed in this experiment.

Reaction Types

Some reactions may be classified as more than one type of reaction; for example, combination and decomposition reactions that involve elemental substances are also oxidation-reduction reactions

Combination Reactions occur when two or more substances (elements or compounds) combine to form one new substance. For example, hydrogen and oxygen gases combine to give water:

\[ 2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l) \]

Decomposition Reactions occur when a compound breaks apart to yield two or more new substances. As an example, potassium chlorate decomposes when heated to yield potassium chloride and oxygen gas. Potassium chlorate is one of the ingredients used in match heads. This reaction occurs as the match ignites, causing the match head to be surrounded by an oxygen atmosphere.
2 KClO₃ (s) → 2 KCl (s) + 3 O₂ (g)

Displacement or Single Replacement Reactions involve the displacement of one element in a compound by another element. Several examples of displacement reactions are given below:

Displacement of one metal by another:

\[
Pb (s) + Cu(NO₃)₂ (aq) \rightarrow Cu (s) + Pb(NO₃)₂ (aq)
\]

Displacement of hydrogen gas from an acid by a metal:

\[
Mg (s) + 2 HCl (aq) \rightarrow MgCl₂ (aq) + H₂ (g)
\]

Displacement of hydrogen gas from water by a metal:

\[
2 K (s) + 2 H₂O (l) \rightarrow 2 KOH (aq) + H₂ (g)
\]

Displacement of one halogen by another:

\[
Cl₂ (aq) + NaBr (aq) \rightarrow NaCl (aq) + Br₂ (aq)
\]

Displacement reactions are always classified as oxidation-reduction reactions. For example, in the first reaction given above, elemental lead is oxidized to lead (II) and copper is reduced from copper (II) to elemental copper. Two electrons are transferred from lead to copper in the process:

\[
Pb (s) \rightarrow Pb^{2+} (aq) + 2 e^- \text{ oxidation of lead}
Cu^{2+} (aq) + 2e^- \rightarrow Cu (s) \text{ reduction of copper}
\]

The ability of one metal to displace another depends on their relative ease of oxidation—a more active metal (one that is more easily oxidized) displaces a less active metal. In the first reaction above, lead is more active than copper. The relative activities of metals can be tabulated in an activity series, ranking the metals by relative ease of oxidation. A metal that displaces hydrogen gas from acid is more active than hydrogen. A metal that displaces hydrogen gas from acid, but not from water, is less active than one that can displace hydrogen from water also. The ease with which a substance is oxidized is quantified as its standard oxidation potential; you will learn more about this in the second semester of General Chemistry.

Exchange or Double Replacement or Metathesis Reactions occur when two compounds that form ions in solution react by switching ion partners. Thus, these reactions have the general form \(AB + CD \rightarrow AD + CB\). One of three conditions must be met for these reactions to occur: the formation of an insoluble ionic compound, observed as a precipitate, the formation of a gas, or the production of water from hydroxide and hydrogen ions (an example of an acid-base neutralization). In each case one of the products results by combining two ions that are removed from the solution by the reaction.

Precipitation reactions occur when aqueous solutions of two ionic compounds are mixed and the ions can combine to make a compound that is not soluble in water (the precipitate). For example, sodium phosphate can be used in an exchange reaction to precipitate calcium ions out of hard water as calcium phosphate, a reaction that is used in some commercial water softeners.

\[
3 CaCl₂ (aq) + 2 Na₃PO₄ (aq) \rightarrow Ca₃(PO₄)₂ (s) + 6 NaCl (aq)
\]

The solubility behavior of the ions that you will be using is summarized in the following table:
Summary of Solubility Rules

<table>
<thead>
<tr>
<th>Solubility Rule</th>
<th>Important Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Group IA, ammonium and nitrate compounds are soluble</td>
<td>None</td>
</tr>
<tr>
<td>Chlorides, Bromides and iodides are soluble.</td>
<td>The halides of silver, lead(II), and mercury(I) are insoluble.</td>
</tr>
<tr>
<td>Sulfates are soluble.</td>
<td>The sulfates of lead(II), barium, and calcium are insoluble.</td>
</tr>
<tr>
<td>Carbonates are insoluble.</td>
<td>The carbonates of Group IA and ammonium are soluble.</td>
</tr>
<tr>
<td>Sulfides are insoluble.</td>
<td>The sulfides of Group IA, ammonium, calcium, and barium are soluble.</td>
</tr>
<tr>
<td>Hydroxides are insoluble.</td>
<td>The hydroxides of Group IA are soluble. Barium hydroxide is soluble, but other Group IIA hydroxides are moderately soluble to insoluble.</td>
</tr>
</tbody>
</table>

Gas forming reactions typically go to completion because one or more of the products are removed from the reaction vessel via the formation of a gas as bubbles that leave the reaction mixture. Gases thus produced include hydrogen sulfide, sulfur dioxide, carbon dioxide and ammonia.

**Hydrogen sulfide**, $H_2S$, by direct combination of acid (source of $H^+$) and sulfide ion:

$$Na_2S(aq) + 2 HCl (aq) \rightarrow H_2S (g) + 2 NaCl (aq)$$

**Sulfur dioxide**, $SO_2$, by decomposition of sulfurous acid initially formed in a reaction between acid and sulfite ion:

$$Na_2SO_3 (s) + 2 HCl (aq) \rightarrow H_2SO_3 (aq) + 2 NaCl (aq)$$

$$H_2SO_3 (aq) \rightarrow H_2O (l) + SO_2 (g)$$

$$Na_2SO_3 (s) + 2 HCl (aq) \rightarrow H_2O (l) + SO_2 (g) + 2 NaCl (aq)$$

**Carbon dioxide**, $CO_2$, by decomposition of carbonic acid initially formed in a reaction between acid and carbonate ion:

$$Na_2CO_3 (s) + 2 HCl (aq) \rightarrow H_2CO_3 (aq) + 2 NaCl (aq)$$

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

$$Na_2CO_3 (s) + 2 HCl (aq) \rightarrow H_2O (l) + CO_2 (g) + 2 NaCl (aq)$$

Ammonia, $NH_3$, from the combination of ammonium and hydroxide ions:

$$NH_4Cl (aq) + NaOH (aq) \rightarrow NaCl (aq) + H_2O (l) + NH_3 (g)$$

**Acid-Base or Neutralization Reactions** involve the transfer of a proton ($H^+$) from the acid to the base:

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

Water is always a product when the base contains hydroxide ion. Some of the gas-forming reactions may also be classified as neutralization reactions.
Procedure

Safety
Be especially cautious with the 6M and 3M acid solutions as they can burn your skin. Do not stare directly at the magnesium when it burns as the light can hurt your eyes. Skin discoloration will result from contact with silver nitrate. Do not touch metals with your hands.

Materials and Equipment
Solids: Mg, CuSO₄•5H₂O, Ca, Cu, Zn, NaHCO₃
Solutions: 6 M HCl, 0.1 M CuSO₄, 0.1 M ZnSO₄, 0.1 M AgNO₃, 0.1 M NaCl, 0.1 M Ni(NO₃)₂, 6 M NaOH, 0.1 M Pb(NO₃)₂, 0.1 M K₂CrO₄, 1 M MH₄NO₃, 6 M H₂SO₄.
Equipment: crucible tongs, one large test tube, two small test tubes, ten small test tubes, test tube holder, test tube rack, 100-mL beaker, red litmus paper, Bunsen Burner
Reagents and Equipment for Instructor Demonstrations: CaO solution (prepared in advance by stockroom), sucrose, 18 M H₂SO₄, distilled water, Na, means of cutting Na and removing from its storage vessel, two 100-mL beakers, straw, red and blue litmus paper, glass stirring rod

Instructions for performing reactions
Perform each of the following reactions except those that are to be demonstrated by your instructor. You use 1 mL of solution in many of the reactions; estimate this by drops (typically 12-15 drops from the reagent bottle dispenser) or by measuring 1 mL once in your graduated cylinder and then transferring it to a test tube to determine how far it fills the test tube. Record your observations on the data page as you do each reaction. Make sure that you observe the results of every reaction even if you didn’t actually mix the chemicals yourself. Write a balanced equation for each reaction. Be sure to include the states of all compounds in your reactions (solid, liquid, aqueous, or gas). If no reaction occurs write the words "no reaction" (or NR) instead of the products in your balanced equation and indicate why you think there was no reaction. Unless otherwise indicated, dispose of all waste in the waste container, or a beaker that you pour into the waste container at the end of the experiment. Do not put metal strips in the sink.

A. Combination Reactions
1. Instructor Demonstration. Pour about 35 mL of a clear saturated solution containing calcium oxide into a 100 mL beaker. Allow the solution to stand for about 15 minutes. Observe. Use a straw to blow bubbles into the solution for a few seconds. Observe the solution again.

2. Hold a small strip of magnesium metal (used in flashbulbs and fireworks) in your crucible tongs and ignite the metal in the hot portion of a burner flame. Don't forget to note the color and composition of the residue left on the tongs.

B. Decomposition Reactions
1. Instructor Demonstration. Do in the fume hood. Fill a 100 mL beaker about one-third full of granulated sugar (sucrose, C₁₂H₂₂O₁₁). Add about 20 mL of concentrated (18 M) sulfuric acid and stir until mixed well. Continue stirring until the mixture darkens. Observe. Do not touch the reaction products or the beaker with your hands; use a stirring rod to guide the solid product that forms.

2. Place a small amount (an amount that will fit on the end of a spatula) of solid copper(II) sulfate pentahydrate in a medium test tube. Use a test tube holder to hold the tube at about a 45° angle and heat in a burner flame for a few minutes, remembering not to point the tube at anyone in the room. Note any changes in the appearance of the solid and anything else
that appears in the test tube. Allow the solid to cool and add a few drops of water. Observe. Dispose of the copper compound in the waste container.

C. Displacement Reactions
Use 1 mL of each solution in a small test tube unless otherwise specified. Use 0.1 M reagents unless otherwise specified. For reactions involving metals use one piece of metal. Do not put the metal pieces in the sink. If no discernable initial change is noted, let the reaction mixture stand for at least five to ten minutes before observing again. Not all of the combinations will yield observable reactions. Repeat the reaction if there is any doubt with respect to whether or not a reaction occurred.

1. Instructor Demonstration. Cautiously add a small piece of sodium metal to water. Test the resulting solution with red and blue litmus papers (red litmus paper will turn blue in the presence of a base; blue litmus will turn red in the presence of an acid).

2. Calcium metal and water (15 mL) in a large test tube.

3. Zinc metal and water

4. Copper metal and 6 M hydrochloric acid

5. Zinc metal and 6 M hydrochloric acid

6. Zinc metal and 0.1 M copper(II) sulfate

7. Copper metal and 0.1 M zinc sulfate

8. Copper metal and 0.1 M silver nitrate

D. Exchange Reactions
Use 1 mL of each solution in a small test tube unless otherwise specified. Use 0.1 M reagents unless otherwise specified. Be sure to mix the solutions well.

1. Silver nitrate and sodium chloride

2. Nickel nitrate and three drops of 6 M sodium hydroxide

3. Lead(II) nitrate and potassium chromate

4. Ammonium nitrate (1 M) and 6 M sodium hydroxide. Warm the test tube gently by passing it back and forth through a burner flame. Hold a strip of moistened red litmus paper so that it is in the tube without coming in contact with the sides of the tube and note any color changes to the paper. Remove the tube from the flame and quickly and cautiously note the smell.

5. Place 5 mL of 6 M hydrochloric acid in a 100 mL beaker. Carefully add several spatulas of solid sodium bicarbonate. Observe.

6. Combine about 5 mL each of 6 M sodium hydroxide and 6 M sulfuric acid in a large test tube. Mix with a stirring rod. Cautiously feel the outside of the test tube. If you cannot detect anything, make sure that you used the correct concentrations of acid and base. **Be very careful with the concentrated sulfuric acid; it is very caustic and can dissolve skin and clothing.**