CHEMICAL REACTIONS

OBJECTIVES

1. To study reactions between ions in aqueous solutions
2. To observe exothermic and endothermic reactions
3. To study oxidation-reduction reactions
4. To practice balancing equations
5. To learn solubility rules

INTRODUCTION

Part I. Chemical reactions of ions in solutions

Ionic compounds that are soluble in water dissolve through a process called dissociation. Dissociation occurs when an ionic compound separates into its component positive and negative ions. For example,

$$NaCl \text{ ------> } Na^{+} + Cl^{-} \qquad (1)$$

Ions in solution are observed to react when an insoluble compound is formed or when a gas is produced. For example,

$$Pb(NO_3)_2(aq) + 2NaCl(aq) \text{ -----> } PbCl_2(s) + 2NaNO_3(aq) \qquad (2)$$

In this reaction a white insoluble compound, $PbCl_2$, is produced. The nitrate and sodium ions do not participate in the reaction (they are spectator ions) so they are normally omitted from the net ionic chemical equation:

$$Pb^{+2} + 2Cl^{-} \text{ -----> } PbCl_2(s) \qquad (3)$$

When sodium bicarbonate, e.g. baking soda, is combined with an acid, a gas is produced

$$HCO_3^{-} + H^{+} \text{ -----> } CO_2(g) + H_2O(l) \qquad (4)$$

When two ionic solutions are mixed and there is not an ion combination that is insoluble or gaseous, no reaction will be observed. For example, no reaction occurs between NaCl and KNO_3 because all ionic combinations are soluble compounds.

Part II. Exothermic and endothermic reactions

Reactions that evolve heat are called exothermic reactions. In an exothermic reaction, the products have less energy than the reactants. When the products have more energy than the reactants, the reaction is endothermic. The additional energy needed for formation of the products is absorbed from the surroundings and can be observed as a cooling of the reaction vessel.

Part III. Oxidation-reduction reactions

When a metal is dissolved in acid, it loses electrons and becomes oxidized. For example,

$$Zn(s) + 2H^{+}(aq) \text{ ------> } Zn^{+2}(aq) + H_2(g) \qquad (5)$$

Oxidation is the loss of electrons. When the Zn atoms lose their electrons, they are gained by the hydrogen ions to produce hydrogen gas. Reduction is the gain of electrons. An oxidation reaction ALWAYS occurs with a corresponding reduction reaction. If an active solid metal is placed in a solution of less active metal ions, an oxidation-reduction (redox) reaction will occur. For example, magnesium is more active than aluminum. If solid magnesium is placed in a solution of aluminum nitrate, the following reaction will occur

$$3Mg(s) + 2Al^{+3}(aq) \text{ ------> } 3Mg^{+2}(aq) + 2Al(s) \qquad (6)$$

Note that the reaction must be balanced such that the number of electrons lost by the magnesium equals the number of electrons gained by the aluminum.

A battery or voltaic cell utilizes a redox reaction to generate electricity (a flow of electrons). One type of cell uses the displacement reaction

$$Zn(s) + Cu^{+2}(aq) \text{ ------> } Zn^{+2}(aq) + Cu(s) \qquad (7)$$

PROCEDURE

Be sure to record your observations for each of the reactions in parts I, II, and III.

Part I. Chemical reactions of ions in solution

A. Reactions that produce solids

Reaction 1: Place 3 drops of $AgNO_3$ solution into a well, then add 3 drops of $CuCl_2$ solution.
Reaction 2: Place 3 drops of Na_2CO_3 solution into a well, then add 3 drops of $Pb(NO_3)_2$ solution.

Reaction 3: Place 3 drops of K_3PO_4 solution into a well, then add 3 drops of $CuCl_2$ solution.
Reaction 4: Place 3 drops of $CuCl_2$ solution into a well, then add 3 drops of Na_2CO_3 solution.

B. Reactions that produce a gas

Reaction 1: Place about 1/4 teaspoon $CaCO_3$ into a well, then add 5 drops of 6 M HCl.

Reaction 2: Place 5 drops of Na_2CO_3 into a well, then add 5 drops of 6M HCl.

Part II. Exothermic and endothermic reactions

Reaction 1: In a small test tube, mix one piece of sodium hydroxide (NaOH) with 2- 5 ml of water. (Be sure to feel the test tube.)

Reaction 2: Place approximately 1/2 teaspoon of ammonium nitrate (NH_4NO_3) into a small test tube. (Use the end of your scoopit.) Add about 5 ml of water.

Part III. Redox reactions

Reaction 1: This reaction must be completed in a hood! Place a piece of copper metal into a small test tube. Add about 1-2 ml of concentrated nitric acid.

Reaction 2: Add 5 drops of $HgNO_3$ to a well of your 12-well microplate. Add 5 drops of NaCl.

CHEMICAL REATIONS NAME _____
LABORATORY REPORT DATE_____

Part I. Chemical reactions of ions in solution

A.

REACTANTS	OBSERVATIONS
$AgNO_3 + CuCl_2$	
$Na_2CO_3 + Pb(NO_3)_2$	
$K_3PO_4 + CuCl_2$	
$CuCl_2 + Na_2CO_3$	

B.

REACTANTS	OBSERVATIONS
$CaCO_3 + HCl$	
$Na_2CO_3 + HCl$	

Part II. Exothermic and endothermic reactions

REACTANTS	OBSERVATIONS
$NaOH + H_2O$	
$NH_4NO_3 + H_2O$	

Part III. Oxidation reduction reactions

REACTANTS	OBSERVATIONS
$Cu + HNO_3$	
$HgNO_3 + NaCl$	

QUESTIONS

Part I.

1. The unbalanced molecular chemical equations for the 4 reactions in Part I A are given below. Write the balanced molecular equations for each. Be sure to indicate the state of each product [(s) for solid, (l) for liquid, (g) for gas, (aq) for aqueous]. You may need to refer to the list of solubility rules in your text (chapter 9).

$AgNO_3 + CuCl_2 \text{ -----> } AgCl + Cu(NO_3)_2$

$Na_2CO_3 + Pb(NO_3)_2 \text{ -----> } PbCO_3 + NaNO_3$

$K_3PO_4 + CuCl_2 \text{ -----> } Cu_3(PO_4)_2 + KCl$

$CuCl_2 + Na_2CO_3 \text{ -----> } CuCO_3 + NaCl$

2. Write balanced reaction(s) for the experiments in Part I B that resulted in gas bubbles. Indicate the gas produced.

3. In Part II, which reaction was exothermic? How did you know? Which was endothermic? How did you know?

Exothermic Reaction:

Why:

Endothermic Reaction:

Why:

4. In Part III, two redox reactions were studied. The first reaction was:

$Cu + 4HNO_3 \text{ -----> } Cu(NO_3)_2 + 2NO_2 + 2H_2O$

The $Cu(NO_3)_2$ gave the solution the green color. NO_2 was the gas produced.

a. Was the copper metal oxidized or reduced?

b. Now that you know what happened to the copper, what do you conclude about the nitrate ion? Was it oxidized or reduced?

5. The second reaction in Part III was

$2HgNO_3 + 2NaCl \text{ ----> } Hg_2Cl_2 + 2NaNO_3$

What is the formula of the white precipitate?

6. Balance the following reactions.

$AgCH_3COO(aq) + NH_4Cl(aq) \text{ -----> } AgCl(s) + NH_4CH_3COO(aq)$

$Na_2CO_3(aq) + NH_4Cl(aq) \text{ -----> } (NH_4)_2CO_3(aq) + NaCl(aq)$

$NaHCO_3(aq) + HCl(aq) \text{ -----> } NaCl(aq) + H_2O(l) + CO_2(g)$

$K_2SO_4(aq) + Pb(NO_3)_2(aq) \text{ -----> } PbSO_4(s) + KNO_3(aq)$

$Na_2S(aq) + Cd(NO_3)_2(aq) \text{ -----> } CdS(s) + NaNO_3(aq)$

7. For the following redox reactions, place a box around the reactant that is being oxidized. Draw a circle around the reactant that is being reduced.

$Mg + H^+ \text{ --------> } Mg^{2+} + H_2$

$Na + Zn^{+2} \text{ -------> } Na^+ + Zn$

$Ag^+ + Cu \text{ ------> } Ag + Cu^{+2}$

$Sn^{+2} + Fe^{+3} \text{ ------> } Sn^{+4} + Fe^{+2}$

$Cl_2 + Br^- \text{ -------> } Cl^- + Br_2$