Experiment 9 Name: ___________________________

DOUBLE DISPLACEMENT REACTIONS

In this experiment, you will observe double displacement reactions and write the corresponding balanced chemical equation and ionic equations.

Double Displacement Reactions

A double displacement reaction is a reaction between two compounds that swap ion components to form two new compounds:

General: AX + BY → AY + BX
Example: 1 $Pb(NO_3)_2$ (aq) + 2 KI (aq) → 1 PbI_2 (s) + 2 KNO_3 (aq)

When a double displacement reaction occurs, one of the following three evidences are observed:

(1) A precipitate forms, when one of the products is an insoluble solid. See Example 1 below.
(2) Heat is given off, when the reaction is between an acid and a base. See Example 2.
(3) Gas produced when one of the products is a gas, such as CO_2. See Example 3.

In this experiment, you will observe these reaction evidences occurring. You will also write the balanced chemical equation and the complete ionic and net ionic equations for the reaction. To write the equations, use the guidelines outlined here.

Writing Ionic Equations for Double Displacement Reactions

1. Write the balanced chemical equation. This equation shows the formulas of the reactants and products.

    a. For both reactants, determine the ions and their charges. Switch cation-anion partners to form the products. Write the formulas of the products based on the ion charges, then balance the chemical equation using coefficients. Writing the formula of the products and balancing the equation are two separate steps. Do not try to balance the equation with subscripts.

> **Note for when the predicted product is H_2CO_3 or H_2SO_3:**
> - The product H_2CO_3 is unstable and spontaneously decomposes into H_2O and CO_2. Write H_2O (l) + CO_2 (g) instead of H_2CO_3 (aq).
> - The product H_2SO_3 is unstable and spontaneously decomposes into H_2O and SO_2. Write H_2O (l) + SO_2 (g) instead of H_2SO_3 (aq).

    b. Designate the phases of the reactants and products.
        (aq) = form aqueous solutions; ionic compound is soluble in water
        (s) = form solid precipitate; ionic compound is insoluble in water
        (l) = liquid, i.e. H_2O
        (g) = gas, i.e. CO_2 and SO_2

To determine whether an ionic compound is soluble or insoluble in water, designated as (aq) or (s), respectively, use the solubility rules as a guideline:

Solubility rules:

Soluble Compounds	**Exceptions**
NO_3^-, ClO_3^-, ClO_4^-, $C_2H_3O_2^-$	
NH_4^+, group I elements	
Cl^-, Br^-, I^-	Ag^+, Hg_2^{2+}, Pb^{2+}
SO_4^{2-}	Sr^{2+}, Ba^{2+}, Hg_2^{2+}, Pb^{2+}

Insoluble Compounds	**Exceptions**
SO_3^{2-}, CrO_4^{2-}, CO_3^{2-}, PO_4^{3-}	NH_4^+, group I elements
OH^-	NH_4^+, group I elements, Ca^{2+}, Sr^{2+}, Ba^{2+}
S^{2-}	NH_4^+, group I elements

2. Write the complete ionic equation. A complete ionic equation shows the actual species present in the solution, that is, with the strong electrolytes dissociated as ions.

    a. Formulas of all compounds designated as (s), (l) or (g) and of molecular compounds are kept intact, they do not dissociate.

    b. For each compound designated (aq), determine if the solution it forms is a strong electrolyte or a weak electrolyte. If it is a strong electrolyte, dissociate the compound into ions, and if it is a weak electrolyte, keep the formula as is.

    Most (aq) compounds are strong electrolytes and are dissociated into ions in solution, but there are exceptions among the compounds of hydrogen (acids) and hydroxide (bases). If the aqueous H^+ or OH^- compound is in the list below, it is a strong acid or base; it is a strong electrolyte and ionize in solution.

    Strong acids: HNO_3, $HClO_4$, $HClO_3$, HCl, HBr, HI, H_2SO_4
    Strong bases: $LiOH$, $NaOH$, KOH, $RbOH$, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$

3. Write the net ionic equation. A net ionic equation shows only the species that participate in the reaction. Ions that appear on both sides of the complete ionic equation, the spectator ions, are omitted.

    Note: The charge of an ion does not change from reactant to the product (unless otherwise told). See Pb^{2+} is Example 1.

Note: Subscripts in the balanced equation become coefficients in the ionic equations. Coefficients in the balanced equation are also coefficients in the ionic equations. If an ion has both a subscript and a coefficient in the balanced equation, the two numbers are multiplied together to become the coefficient in the ionic equations. (See Example 4)

Example 1: $Pb(NO_3)_2$ + KI

1. Balanced chemical equation – The ions are Pb^{2+} and NO_3^- in $Pb(NO_3)^2$ ad K^+ and I^- in KI. When the ions switch partners, the products are PbI_2 and KNO_3. Coefficients are then used to balance the equation. Referring to the solubility rules, $Pb(NO_3)_2$, KI and KNO_3 are soluble in water and PbI_2 is insoluble. The formation of a yellow precipitate, PbI_2, is observed when this reaction takes place.

$$1\ Pb(NO_3)_2\ (aq) + 2\ KI\ (aq) \rightarrow 1\ PbI_2\ (s) + 2\ KNO_3\ (aq)$$

2. Complete ionic equation – The compounds designated (aq) in this reaction all form strongly electrolytic solutions and thus dissociate into ions.

$$1\ Pb^{2+}\ (aq) + 2\ NO_3^-\ (aq) + 2\ K^+\ (aq) + 2\ I^-\ (aq) \rightarrow 1\ PbI_2\ (s) + 2\ K^+\ (aq) + 2\ NO_3^-\ (aq)$$

3. Net ionic equation – 2 NO_3^- (aq) and 2 K^+ (aq), the spectator ions, appear on both sides of the complete ionic equation and are omitted.

$$1\ Pb^{2+}\ (aq) + 2\ I^-\ (aq) \rightarrow 1\ PbI_2\ (s)$$

Example 2: $HClO_4$ + NaOH

1. Balanced chemical equation – When this reaction takes place, heat is given off. This indicates that one of the products is water, as what occurs during an acid-base (neutralization) reaction.

$$1\ HClO_4\ (aq) + 1\ NaOH\ (aq) \rightarrow 1\ NaClO_4\ (aq) + 1\ \mathbf{HOH\ (l)}$$

Note: HOH should be rewritten as H_2O since all atoms of the same type are grouped together in a chemical formula.

$$1\ HClO_4\ (aq) + 1\ NaOH\ (aq) \rightarrow 1\ NaClO_4\ (aq) + 1\ \mathbf{H_2O\ (l)}$$

2. Complete ionic equation – $HClO_4$, NaOH, and $NaClO_4$ break up into ions in solution.

$$1\ H^+\ (aq) + 1\ ClO_4^-\ (aq) + 1\ Na^+\ (aq) + 1\ OH^-\ (aq) \rightarrow 1\ Na^+\ (aq) + 1\ ClO_4^-\ (aq) + 1\ H_2O\ (l)$$

3. Net ionic equation – Na^+ and ClO_4^- are the spectator ions and cancel out.

$$1\ H^+\ (aq) + 1\ OH^-\ (aq) \rightarrow +\ 1\ H_2O\ (l)$$

Example 3: Li_2CO_3 + HBr

1. Balanced chemical equation – The double displacement products are LiBr and H_2CO_3, but H_2CO_3 is unstable and decomposes to H_2O and CO_2. In this reaction, gas bubbles form due to CO_2 production.

$$1\ Li_2CO_3\ (aq) + 2\ HBr\ (aq) \rightarrow 2\ LiBr\ (aq) + 1\ H_2O\ (l) + 1\ CO_2\ (g)$$

2. Complete ionic equation – Li_2CO_3. HBr and LiBr dissociate into ions. H_2O and CO_2 are molecular compounds and do not ionize.

$$2\ Li^+\ (aq) + 1\ CO_3^{2-}\ (aq) + 2\ H^+\ (aq) + 2\ Br^-\ (aq) \rightarrow 2\ Li^+\ (aq) + 2\ Br^-\ (aq) + 1\ H_2O\ (l) + 1\ CO_2\ (g)$$

3. Net ionic equation – Li^+ and Br^- are the spectator ions and cancel out.

$$1\ CO_3^{2-}\ (aq) + 2\ H^+\ (aq) \rightarrow 1\ H_2O\ (l) + 1\ CO_2\ (g)$$

Example 4: $FeCl_3$ + $Ba(OH)_2$

1. Balanced chemical equation –.This reaction results in the formation of a precipitate ($Fe(OH)_3$).

$$2\ FeCl_3\ (aq) + 3\ Ba(OH)_2\ (aq) \rightarrow 2\ Fe(OH)_3\ (s) + 3\ BaCl_2\ (aq)$$

2. Complete ionic equation – $FeCl_3$, $Ba(OH)_2$, and $BaCl_2$ break up into ions in solution. Note that Cl^- has a coefficient of 6, because 2 x 3 = 6. Each molecule of $FeCl_3$ contains 3 Cl^- and there are 2 $FeCl_3$ in the balanced equation, hence 6 Cl^- ions.

$$2\ Fe^{3+}\ (aq) + 6\ Cl^-\ (aq) + 3\ Ba^{2+}\ (aq) + 6\ OH^-\ (aq) \rightarrow 2\ Fe(OH)_3\ (s) + 3\ Ba^{2+}\ (aq) + 6\ Cl^-\ (aq)$$

3. Net ionic equation – Cl- and Ba^{2+} are the spectator ions and cancel out.

$$2\ Fe^{3+}\ (aq) + 6\ OH^-\ (aq) \rightarrow +\ 2\ Fe(OH)_3\ (s)$$

PROCEDURE

For each reaction, write the balanced chemical equation and the complete and net ionic equations. Make sure that all equations are balanced, and all phase designations are indicated. Write the charges for ions in the ionic equations and check that the overall charges are balanced on both sides of the equations.

The equations should correspond to your observation; for example, if you observe bubble formation, your equation shows that one of the products is a gas.

When there are no observable changes, write "no reaction" under the observations. Write the balanced chemical equation, and then the complete ionic equation, which shows all the ions on the reactant and product sides cancel each other out, resulting to no net ionic equation. Write "no reaction" under net ionic equation.

For reactions 1 through 11, mix equal volumes (~10 drops each) of the two solutions in a test tube. Record your observations; remember that you are watching for formation of precipitate (note the color), release of heat or evolution of gas. **Reaction 12 must be done in the hood**. Follow the special instructions for reaction 12.

IMPORTANT: Dispose wastes as soon as you finish a reaction or part so you can keep track of the test tube contents and place them in their designated containers.

1. 0.1 M sodium chloride + 0.1 M potassium nitrate
2. 0.1 M sodium chloride + 0.2 M silver nitrate
3. 0.1 M sodium carbonate + 6 M hydrochloric acid
4. 10% sodium hydroxide + 6 M hydrochloric acid
5. 0.1 M barium chloride + 3 M sulfuric acid
6. 6 M ammonium hydroxide + 3 M sulfuric acid
7. 0.1 M copper(II) sulfate + 0.1 M zinc nitrate
8. 0.1 M sodium carbonate + 0.1 M calcium chloride
9. 0.1 M copper(II) sulfate + 0.1 M ammonium chloride
10. 10% sodium hydroxide + 3 M nitric acid
11. 0.1 M iron(III) chloride + 6 M ammonium hydroxide
12. **This reaction must be done in the hood.** Place ~0.25 g solid sodium sulfite in a test tube. Add 20 drops of distilled water. Then add ~20 drops of 6 M hydrochloric acid.

Keep in mind that not all reactions are instantaneous. Allow time for the reaction to occur.

CLEAN-UP

- Dispose of wastes in designated containers in the front hood.
- Wash all test tubes. Shake off excess water and return to the test tube racks on your work station.

Experiment 9 Name: ______________________

Name: ______________________________ **Date:** __________________
Partner's Name: _________________________

DOUBLE DISPLACEMENT REACTIONS

1. NaCl (aq) + KNO_3 (aq)
Observations
Balanced Chemical Equation
Complete Ionic Equation
Net Ionic Equation
2. NaCl (aq) + $AgNO_3$ (aq)
Observations
Balanced Chemical Equation
Complete Ionic Equation
Net Ionic Equation

3. Na_2CO_3 (aq) + HCl (aq)
Observations
Balanced Chemical Equation
Complete Ionic Equation
Net Ionic Equation
4. NaOH (aq) + HCl (aq)
Observations
Balanced Chemical Equation
Complete Ionic Equation
Net Ionic Equation

5. $BaCl_2$ (aq) + H_2SO_4 (aq)
Observations
Balanced Chemical Equation
Complete Ionic Equation
Net Ionic Equation
6. NH_4OH (aq) + H_2SO_4 (aq)
Observations
Balanced Chemical Equation
Complete Ionic Equation
Net Ionic Equation

7. $CuSO_4$ (aq) + $Zn(NO_3)_2$ (aq)
Observations
Balanced Chemical Equation
Complete Ionic Equation
Net Ionic Equation

8. Na_2CO_3 (aq) + $CaCl_2$ (aq)
Observations
Balanced Chemical Equation
Complete Ionic Equation
Net Ionic Equation

9. $CuSO_4$ (aq) + NH_4Cl (aq)
Observations
Balanced Chemical Equation
Complete Ionic Equation
Net Ionic Equation
10. NaOH (aq) + HNO_3 (aq)
Observations
Balanced Chemical Equation
Complete Ionic Equation
Net Ionic Equation

11. $FeCl_3$ (aq) + NH_4OH (aq)
Observations
Balanced Chemical Equation
Complete Ionic Equation
Net Ionic Equation

12. Na_2SO_3 (aq) + HCl (aq) **DO IN THE FUME HOOD!**
Observations
Balanced Chemical Equation
Complete Ionic Equation
Net Ionic Equation