Experiment 10
Double Displacement Reactions

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Five signs a reaction occurred.
2. Write complete Double Displacement chemical reactions
   (a) Balance charges.
   (b) Balance atoms.
   (c) Include states when known.
3. Recognize when no reaction (NR) occurs.
4. Use a solubility table to determine the state of compounds.

Discussion

Double displacement reactions are among the most common of the simple chemical reactions to study and understand. We will explore the driving forces behind the chemical reactions, and use observations made about individual reactions to write complete chemical equations.

In examining the characteristic chemical properties of several compounds we will recall that there are several readily apparent signs that a chemical reaction between two solutions has occurred. Those signs were:

1. Color changes, that are not the result of simple dilution.
2. Evolution of a gas forming many bubbles rapidly.
3. Formation of a precipitate, normally a solid which will settle to the bottom of a test tube, but occasionally a finely divided precipitate that does not settle and often appears as only a milky color.
4. Disappearance of a precipitate.
5. Endothermic or Exothermic reactions (creation of heat or cold).

Each of these signs will be used in this experiment to determine if a chemical reaction has occurred. For those experiments in which a reaction occurred we will write a balanced chemical reaction, including the states of products based on our observations.

Chemical reactions occur for a number of reasons. A detailed discussion is beyond the scope of this lab, however a brief discussion will provide insight into why each of the observable signs of a chemical reaction occur. **Enthalpy** is a measure of the energy flow in a reaction, while **Entropy** is a measure of the randomness or disorder in a system. Spontaneous chemical reactions seek to decrease the enthalpy of the products and increase the entropy of the products. One or both of these conditions

must be met for the reaction to occur spontaneously. It is the interplay of these two mechanisms that will drive the chemical reactions we see.

When examining the physical states of the reactants and products entropy increase as we go from a solid to a liquid to a gas, because the motion (disorder) of the particles is increased. Enthalpy is more difficult to predict and will be covered in later lectures or experiments. As a general rule the formation small stable molecules (listed in Table 10.1 below or the formation of water (H_2O) is favorable and a reaction will occur. One can also use solubility tables to determine which solids are more stable and tend to form precipitates in solution.

The formation of precipitates can drive a reaction because the products are more stable and lower energy than the reactants. The production of gases is both an increase in the disorder in a system, and often results in the formation of small stable molecules. An exothermic reaction is one in which heat or energy is given off indicating that the products are lower in energy than the reactants. Color changes are a good sign that a chemical reaction has occurred, but are not directly tied to changes in enthalpy or entropy, and instead occur due to structural changes in how the atoms in the solution interact with light.

Identifying precipitates and gases

Recognizing precipitates and gases while performing a chemical reaction is very obvious, however when writing chemical reactions on paper, we need several aids to help us identify them.

There are several common gases it is worth memorizing the state of shown in Table 10.1

Common Gases
Elemental (11) - H_2, N_2, O_2, F_2 and Cl_2 + Noble Gases
Molecular (8) - CO, CO_2, NO, NO_2, N_2O_4, SO_2, SO_3, and NH_3

Table 10.1: Common gasses formed in chemical reactions.

Gases can also be formed by the decomposition of several common compounds formed in double displacement reactions. There are three common reactions that we will encounter and need to recognize given below. The formation of any of these compounds will result in the formation of a gas, indicating a chemical reaction has occurred.

$$\mathrm{H_2CO_3(d)} \longrightarrow \mathrm{H_2O(l) + CO_2(g)} \tag{9}$$

$$\mathrm{H_2SO_3(d)} \longrightarrow \mathrm{H_2O(l) + SO_2(g)} \tag{10}$$

$$\mathrm{NH_4OH(d)} \longrightarrow \mathrm{H_2O(l) + NH_3(g)} \tag{11}$$

Precipitates can be determined by consulting solubility tables. These are tables of common cations and anions arranged in columns and rows. The intersection of which gives the information about solubility. A soluble compound is one that dissolves in water, while insoluble compounds will not dissolve and thus form precipitates. A solubility table is included at the end of this experiment. Thus, if we examine the solubility of NaCl for example we note that it is labeled “(aq)” in the table indicating that no precipitate occurs, and thus no reaction is likely to have occurred. An example of when a precipitate occurs is the

reaction of Ag + Cl to form AgCl which is labeled as "I" indicating that it would form a precipitate and a reaction would occur.

A. Exothermic and Endothermic Reactions

Recognizing temperature while performing a chemical reaction is very obvious, however when writing chemical reactions on paper, we need several aids to help us identify them. In this experiment we will only examine exothermic reactions, the classic example being the reaction of an acid and a base. The general reaction is shown below.

$$\text{Acid} + \text{Base} \longrightarrow \text{Ionic Salt} + \text{H}_2\text{O(l)} + \text{Heat} \tag{12}$$

Table 10.2 lists the most common acids and bases. In order to write the chemical reactions it will be necessary to recognize them in chemical reactions.

The heat in an acid/base reaction occurs because of the formation of a slightly ionized compound (in this case water). Thus, we can add an additional sign that a reaction has occurred which is whenever water is formed in a chemical reaction. The heat is produced because water is a highly stable molecule, and when formed is lower in energy than the reactants, resulting in the release of heat. Other examples of slightly ionized compounds formed in double displacement reactions are acetic acid ($HC_2H_3O_2$), oxalic acid ($H_2C_2O_4$), and phosphoric acid (H_3PO_4).

Acids (start with H)	HF, HCl, HBr, HI, H_3PO_4, H_2SO_4, HNO_3, H_2CO_3, $HC_2H_3O_2$, $HClO_4$
Bases (contain OH)	NaOH, KOH, $Ba(OH)_2$, $Ca(OH)_2$, NH_4OH

Table 10.2: Common Acids and Bases.

Double Displacement Reactions

In the double displacement reactions in this experiment two aqueous solutions each containing an ionic compound will be mixed. The general reaction is given by:

$$\text{AB} + \text{CD} \longrightarrow \text{AD} + \text{CB} \tag{13}$$

Each ionic compound (AB and CD) when dissolved in water will be present as a cation and anion pair. This implies that when the solutions are mixed we will have a mixture containing A^+, B^-, C^+, and D^- cations and anions. As the ions interact with each other there are six possible combinations of ions that could result in a reaction. The six outcomes are shown in Table 10.3.

Double displacement reactions are named because the reaction can be viewed as one cation displacing the other cation from the molecule and forming a new compound. These reactions are also often referred to as Double Replacement reactions since we can view one cation as replacing the other to form a new compound. A third way of viewing these reactions is illustrated below in Figure 10.1 where we can imagine that the cations simply swap places and each of the new combinations has the potential to cause a reaction to occur.

Combination	Result
A^+B^- C^+D^-	The ions recombine into the starting reactants. This can not result in a reaction because no change occurs.
A^+C^+ B^-D^-	This can not result in a reaction because like charges repel each other. No exchange of electrons can occur to form a compound.
A^+D^- C^+B^-	This may result in a reaction if a precipitate or gas is formed.

Table 10.3: Possible outcomes of a double displacement reaction.

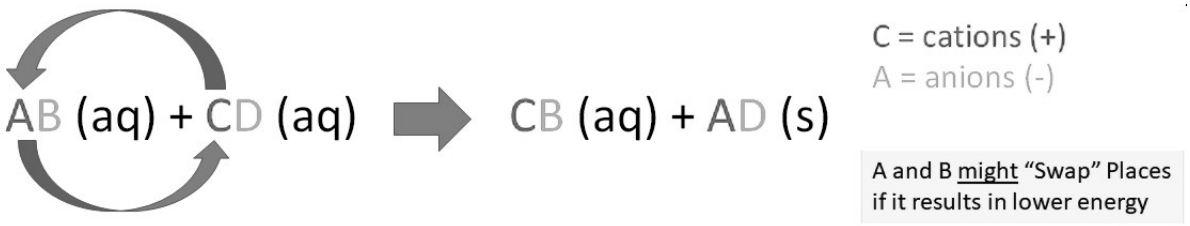

Figure 10.1: Double displacement reactions occur if the products formed by swapping cations are lower in energy than the reactants. credit: author

Example 1:

A solution of sodium chloride and lead (II) nitrate are mixed. A white precipitate is formed. Write the complete chemical reaction.

The precipitate is definite evidence that a chemical reaction occurred. We swap the sodium and lead cations to form new compounds, and checking the solubility tables we note that $PbCl_2$ is listed as (I), indicating it forms a solid precipitate, while $NaNO_3$ is listed as aqueous. We can then write:

$$\underline{2}\,NaCl(aq) + \underline{1}\,Pb(NO_3)_2(aq) \longrightarrow \underline{1}\,PbCl_2(s) + \underline{2}\,NaNO_3(aq) \tag{14}$$

Example 2:

A solution of sodium carbonate and sulfuric acid are mixed, resulting in the formation of a gas. Write the complete chemical reaction.

$$\underline{1}\,Na_2CO_3(aq) + \underline{1}\,H_2SO_4(aq) \longrightarrow \underline{1}\,Na_2SO_4(aq) + \underline{1}\,H_2CO_3(aq) \tag{15}$$

Examining the solubility tables indicates that neither product forms a precipitate. There is an acid in the reaction, however there is no base for it to react with, the experimenter did not note any temperature change. Gas was observed indicating that a chemical reaction must have occurred, and we must recall that H_2CO_3 readily decomposes into water and carbon dioxide. The final chemical reaction would then be:

$$\underline{1}\,Na_2CO_3(aq) + \underline{1}\,H_2SO_4(aq) \longrightarrow \underline{1}\,Na_2SO_4(aq) + \underline{1}\,H_2O(l) + \underline{1}\,CO_2(g) \tag{16}$$

Example 3:

A solution of magnesium hydroxide and hydrochloric acid are mixed, no visible sign of a reaction occurred but the test tube feels warm to the touch. The double displacement reaction is written as:

$$\_1\_\,\mathrm{Mg(OH)_2(aq)} + \_2\_\,\mathrm{HCl(aq)} \longrightarrow \_1\_\,\mathrm{MgCl_2(aq)} + \_2\_\,\mathrm{H_2O(l)} \tag{17}$$

Neither product is a precipitate or gas, however, we should recognize this as an acid/base reaction. Additionally we could note that water was formed in the reaction, indicating a reaction occurred. Both acid/base reactions and the formation of water result in the release of heat. We can write the final reaction then as:

$$\_1\_\,\mathrm{Mg(OH)_2(aq)} + \_2\_\,\mathrm{HCl(aq)} \longrightarrow \_1\_\,\mathrm{MgCl_2(a_1)} + \_2\_\,\mathrm{H_2O(l)} + \mathrm{Heat} \tag{18}$$

Example 4:

A solution of sodium chloride and potassium nitrate are mixed. Write the complete chemical reaction. No change is observed in the test tube.

$$\_\_\mathrm{NaCl(aq)} + \_\_\mathrm{KNO_3(aq)} \longrightarrow \_\_\mathrm{KCl(aq)} + \_\_\mathrm{NaNO_3(aq)} \tag{19}$$

We write the initial reaction by simply swapping the sodium and potassium cations. We then need to check the solubility tables for KCl and $NaNO_3$, from which we note both compounds form aqueous compounds. Neither compound is a gas or decomposes and the reactants are not acids and bases, thus we conclude that no reaction (NR) occurred and would write:

$$\_\_\mathrm{NaCl(aq)} + \_\_\mathrm{KCl(aq)} \longrightarrow \mathrm{NR} \tag{20}$$

Procedure

Obtain 12 medium sized test tubes and a test tube rack. Perform the following 12 reactions as described below. Each part of the experiment involves mixing equal volumes of solution (about 1 mL or 20 drop) in a standard sized test tube. If you are unsure of your observation you may add another milliliter of each solution. For each reaction record your observations at the time of mixing. Feel each tube to determine if heat was evolved. The following terminology should be used:

Observation	**Notation**
Precipitate	ppt - color
Gas	(g)
Heat	Heat
No reaction	NR

For each reaction write the complete chemical equation. Be sure to balance each product formed, and the overall chemical equation. Include the states of the reactants and products when known, if the state is not known leave empty bracket ( ). If the reaction is exothermic add “heat” as a product. If there is no evidence of a reaction write cross out the products and write “No Reaction” or “NR” as the products.

All reactions written should reflect observations (and vice versa), consult with your instructor about any discrepancies.

Perform the following twelve chemical reactions in any order.

1. Mix 0.1 M sodium chloride and 0.1 M potassium nitrate solutions.
2. Mix 0.1 M sodium chloride and 0.1 M silver nitrate solutions.
3. Mix 0.5 M sodium carbonate and concentrated (6 M) hydrochloric acid solutions.
4. Mix 20% sodium hydroxide and 6 M hydrochloric acid solutions.
5. Mix 0.1 M barium chloride and 9 M sulfuric acid solutions.
6. Mix 6 M ammonium hydroxide and 3 M sulfuric acid solutions.
7. Mix 0.1 M copper (II) sulfate and 0.1 M zinc nitrate solutions.
8. Mix 0.1 M sodium carbonate and 0.1 M calcium chloride solutions.
9. Mix 0.1 M copper (II) sulfate and 0.1 M ammonium chloride solutions.
10. Mix 20% sodium hydroxide and 6 M nitric acid solutions.
11. Mix 0.1 M iron (III) chloride and 6 M ammonium hydroxide solutions.
12. In the hood add 1 g of sodium sulfite to a test tube. Then add 2 mL of water and swirl vigorously to dissolve the solid (some small amount of solid may remain, this is acceptable). This produces an aqueous solution of sodium sulfite which we will consider as the reactant. Slowly (drop by drop) add 1 mL of 18 M sulfuric acid to the sodium sulfite solution.

 Dispose of mixtures 2, 5, 7, 9 in the “Experiment 6 Heavy Metal Waste” container. Dispose of all other mixtures in the sink and flush with water.

Name: _____ Class: _____ Date: _____

Results

For each reaction write the observed results from lab. Then write the complete chemical reaction. Be sure to balance each reactant and product compounds charges, and the overall chemical equation. Include the states of the reactants and products when known, if the state is not known leave empty bracket ( ). If the reaction is exothermic add "+ heat" as a product. If there is no evidence of a reaction cross out the products and write "No Reaction" or "NR" as the products. All reactions written should reflect observations (and vice versa), consult with your instructor about any discrepancies.

#	Observations	Complete Chemical Reaction
1.		
2.		
3.		
4.		
5.		
6.		
7.		
8.		
9.		
10.		
11.		
12.		

Table 10.4: Results

Post Lab Questions

1. Write the complete chemical reaction for the decomposition of sulfurous acid (H_2SO_3).

2. Using the criteria for double displacement reactions occurring, solubility tables, and the knowledge gained in this experiment predict whether a double displacement reaction will occur in each example below. If the reaction will occur complete and balance each compound, and equation, properly indicate the state of gases and precipitates, and indicate if the reaction is exothermic by including "heat" as a product. If no reaction will occur write "No reaction" or "NR" as the products. All reactants are assumed to be aqueous.

   (a) $\_\_\ K_2S(aq) + \_\_\ CuSO_4(aq) \longrightarrow$

   (b) $\_\_\ NH_4OH(aq) + \_\_\ H_2C_2O_4(aq) \longrightarrow$

   (c) $\_\_\ KOH(aq) + \_\_\ NH_4Cl(aq) \longrightarrow$

   (d) $\_\_\ Na_2CrO_4(aq) + \_\_\ Pb(C_2H_3O_2)_2(aq) \longrightarrow$

   (e) $\_\_(NH_4)_2SO_4(aq) + \_\_NaI(aq) \longrightarrow$

   (f) $\_\_FeCl_3(aq) + \_\_\ NaOH(aq) \longrightarrow$

   (g) $\_\_KC_2H_3O_2(aq) + \_\_CoSO_4(aq) \longrightarrow$

   (h) $\_\_\ Na_2CO_3(aq) + \_\_\ HNO_3(aq) \longrightarrow$

   (i) $\_\_\ ZnBr_2(aq) + \_\_\ K_3PO_4(aq) \longrightarrow$

   (j) $\_\_\ KOH(aq) + \_\_\ H_3PO_4(aq) \longrightarrow$

Name: _____ Class: _____ Date: _____

Prelab Questions

1. What are the 4 signs that a double displacement reaction has occurred in **LAB**.

2. Complete the table below listing all of the chemical compounds used in lab.

Name	**Formula**
Sodium Chloride	
Potassium Nitrate	
Silver Nitrate	
Sodium Carbonate	
Hydrochloric Acid	
Nitric Acid	
Sodium Sulfite	
	$BaCl_2$
	H_2SO_4
	NH_4OH
	$CuSO_4$
	$Zn(NO_3)_2$
	$CaCl_2$
	$FeCl_3$