Experiment 9 - Double Displacement Reactions

A double displacement reaction involves two ionic compounds that are dissolved in water. In a double displacement reaction, it appears as though the ions are "trading places," as in the following hypothetical reaction:

$$AB_{(aq)} + CD_{(aq)} \rightarrow AD + CB$$

Where AB exists as A^+ and B^- ions in solution and CD exists as C^+ and D^- ions in solution. As the ions come in contact with each other, there are six possible combinations that might conceivably cause a chemical reaction. Two of these combinations are the meeting of ions of like charge; that is, $A^+ + C^+$ and $B^- + D^-$. But since like charges repel, no reaction will occur with these combinations. Two other possible combinations are those of the original two compounds; that is, $A^+ + B^-$ and $C^+ + D^-$. Since we originally had a solution containing each of these pairs of ions, they can mutually exist in the same solution; therefore they do not recombine. Thus the two possibilities for chemical reaction are the combination of each of the positive ions with the negative ion of the other compound; that is, $A^+ + D^-$ and $C^+ + B^-$.

In summary, when the solutions are mixed, these ions can all come into contact with each other, and new products could be formed. If new products are to be formed, there is only one possible combination of products: since like charges repel each other, we cannot have new compounds containing two negative ions or two positive ions. The only other possible <u>new</u> combination comes from the positive and negative ions of the two compounds switching places.

There are three types of equations that can be written for reactions that involve ions in solution. The first type is just the regular, overall equation. An example of an overall equation for a double displacement reaction is:

$$Na_2CO_{3\,(aq)} + 2\ AgNO_{3\,(aq)} \rightarrow 2\ NaNO_{3\,(aq)} + Ag_2CO_{3\,(s)}$$

In this type of equation, the complete formulas are shown along with the appropriate state symbols. The formulas of the products are obtained as follows:

1. Determine what ions are present in the reactants, without worrying about how many of each there are. Include their charges: $Na^+_{(aq)}$, $CO_3^{2-}{}_{(aq)}$, $Ag^+_{(aq)}$, and $NO_3^-{}_{(aq)}$.

2. Switch the ions so that the cation of the first compound is paired with the anion of the second compound, and vice versa. In this case, the $Na^+_{(aq}$ is paired with the $NO_3^-{}_{(aq)}$, and the $Ag^+_{(aq}$ is paired with the $CO_3^{2-}{}_{(aq)}$.
3. Determine the correct formulas of the products, keeping in mind the charges on each of the ions and remembering that the overall formula of an ionic compound must have no net charge. In this case, since sodium ion has a charge of +1 and nitrate ion has a charge of –1, the correct formula of the compound sodium nitrate contains one of each ion: $NaNO_3$. For the other compound, silver ion has a charge of +1 and carbonate ion has a charge of –2. Therefore, two silver ions are needed to balance out the charge on one carbonate ion, and the formula of the compound is Ag_2CO_3.
4. Write the formulas of the products, and check the solubility rules to determine the appropriate state symbol to include next to the products. In this case, according to the solubility rules, sodium nitrate is soluble, but silver carbonate is insoluble.
5. Balance the equation.

Another type of chemical equation, the **complete ionic equation**, shows all substances in the reaction in their predominant form in solution. Since soluble ionic compounds consist of separated ions in solution, they are shown in the equation as separated ions. Likewise, any substances which exist mostly as molecules are shown as molecules in the equation. Insoluble ionic compounds are <u>not</u> shown ionized, since their ions are not actually separated in the solution. The complete ionic equation for the above reaction would be:

$$2\ Na^+_{(aq)} + CO_3^{2-}{}_{(aq)} + 2\ Ag^+_{(aq)} + 2\ NO_3^-{}_{(aq)} \rightarrow 2\ Na^+_{(aq)} + 2\ NO_3^-{}_{(aq)} + Ag_2CO_{3(s)}$$

A third type of equation is called a **net ionic equation**. In a net ionic equation, only those species (ions or molecules) that actually change are shown. All of the **spectator ions** (those ions that are present in the solution but which are not reacting) are not written. In the above reaction, the spectator ions are sodium ion and nitrate ion: they appear on both sides of the reaction unchanged. The net ionic equation for the above reaction is therefore:

$$CO_3^{2-}{}_{(aq)} + 2\ Ag^+{}_{(aq)} \rightarrow Ag_2CO_{3(s)}$$

A net ionic equation is not really a complete equation, since it does not give complete formulas; it is nevertheless quite useful, since it focuses attention on the main event. In the present case, the equation says the Ag^+ ion (from an unspecified source) combines with a CO_3^{2-} ion (also from an unspecified source) to form a precipitate of Ag_2CO_3. This will happen any time these two ions are put together in the same solution.

In each part of this experiment two aqueous solutions, each containing positive and negative ions, will be mixed in a test tube. You will observe whether or not a reaction occurs in each case, and predict the products. You will also practice writing different types of equations for reactions. Let us look at some examples.

Example 1. When solutions of sodium chloride and silver nitrate are mixed, the equation for the hypothetical double displacement reaction is:

$$NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$$

(The formulas of the products are obtained by switching the ions.) A white precipitate is produced when these solutions are mixed. This precipitate is definite evidence of a chemical reaction. One of the two products, sodium nitrate ($NaNO_3$) or silver chloride (AgCl) is insoluble. Although the precipitate could be identified by further chemical testing, we can instead look at the solubility table to find that sodium nitrate is soluble but silver chloride is insoluble. We may then conclude that the precipitate is silver chloride and indicate this in the equation with the state symbol (s), which stands for "solid". Thus the overall equation should read:

$$NaCl_{(aq)} + AgNO_{3\,(aq)} \rightarrow NaNO_{3\,(aq)} + AgCl_{(s)}$$

The complete ionic equation for the above reaction would be:

$$Na^+_{(aq)} + Cl^-_{(aq)} + Ag^+_{(aq)} + NO_3^-{}_{(aq)} \rightarrow Na^+_{(aq)} + NO_3^-{}_{(aq)} + AgCl_{(s)}$$

The net ionic equation for the above reaction (canceling out the spectator ions sodium ion and nitrate ion) is:

$$Cl^-_{(aq)} + Ag^+_{(aq)} \rightarrow AgCl_{(s)}$$

Example 2. When solutions of sodium chloride and potassium nitrate are mixed, the equation for the hypothetical double displacement reaction is

$$NaCl + KNO_3 \rightarrow KCl + NaNO_3$$

(The formulas of the products are obtained by switching the ions.) We get the hypothetical products by simply combining each positive ion with the other negative ion. But has there been a reaction? When we do the experiment we see no evidence of reaction. There is no precipitate formed, no gas evolved, and no obvious temperature change. Thus we must conclude that no reaction occurred. Both hypothetical products are soluble salts, so the ions are still present in solution. We can say that we simply have a solution of the four kinds of ions, Na^+, Cl^-, K^+, and NO_3^-.

The situation is best expressed by changing the equation to

$$NaCl + KNO_3 \rightarrow \text{No reaction}$$

The complete ionic equation for this reaction is:

$$Na^+_{(aq)} + Cl^-_{(aq)} + K^+_{(aq)} + NO_3^-{}_{(aq)} \rightarrow K^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + NO_3^-{}_{(aq)}$$

There is no net ionic equation for this reaction. Since all of the ions are spectator ions, they all cancel out.

Example 3. When solutions of sodium carbonate and hydrochloric acid are mixed, the equation for the hypothetical double displacement reaction is:

$$Na_2CO_3 + 2\ HCl \rightarrow 2\ NaCl + H_2CO_3$$

Bubbles of a colorless gas are evolved when these solutions are mixed. Although this gas is evidence of a chemical reaction, neither of the indicated products is a gas. But carbonic acid, H_2CO_3, is an unstable compound and readily decomposes into carbon dioxide and water.

$$H_2CO_3 \rightarrow H_2O + CO_{2\ (g)}$$

Therefore, CO_2 and H_2O are the products that should be written in the equation. The original equation then becomes

$$Na_2CO_{3\ (aq)} + 2\ HCl_{(aq)} \rightarrow 2\ NaCl_{(aq)} + H_2O_{(l)} + CO_{2\ (g)}$$

The complete ionic equation for this reaction is:

$$2\ Na^+_{(aq)} + CO_3^{2-}{}_{(aq)} + 2\ H^+_{(aq)} + 2\ Cl^-_{(aq)} \rightarrow 2\ Na^+_{(aq)} + 2\ Cl^-_{(aq)} + H_2O_{(l)} + CO_{2\ (g)}$$

The net ionic equation reaction (canceling out the spectator ions sodium ion and chloride ion) is:

$$CO_3^{2-}{}_{(aq)} + 2\ H^+_{(aq)} \rightarrow H_2O_{(l)} + CO_{2\ (g)}$$

Examples of other substances that decompose to form gases are sulfurous acid (H_2SO_3) and ammonium hydroxide (NH_4OH):

$$H_2SO_{3\ (aq)} \rightarrow H_2O_{(l)} + SO_{2\ (g)}$$

$$NH_4OH_{(aq)} \rightarrow H_2O_{(l)} + NH_{3\ (g)}$$

Example 4. When solutions of sodium hydroxide and hydrochloric acid are mixed, the equation for the hypothetical double displacement reaction is:

$$NaOH + HCl \rightarrow NaCl + H_2O$$

The mixture of these solutions produces no visible evidence of reaction, but on touching the test tube we notice that it feels warm. The evolution of heat is evidence of a chemical reaction. This example and Example 2 appear similar because there is no visible evidence of reaction. However, the difference is very important. In Example 2 all four ions are still uncombined. In the present example the hydrogen ions (H^+) and the hydroxide ions (OH^-) are no longer free in solution but have combined to form water. The reaction of H^+ (an acid) and OH^- (a base) is called **neutralization**. The formation of the slightly ionized compound (water) causes the reaction to occur and is the source of the heat given off. The overall balanced equation is:

$$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

Since HCl , NaOH, and NaCl are all strong electrolytes, they are all completely ionized in solution. Water, however, consists mostly of un-ionized water molecules, so it is not shown as separated ions. The complete ionic equation is:

$$Na^+_{(aq)} + OH^-_{(aq)} + H^+_{(aq)} + Cl^-_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$$

The net ionic equation (canceling out the spectator ions sodium ion and chloride ion) is:

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$

Water is the most common slightly ionized substance formed in double displacement reactions; other examples are acetic acid ($HC_2H_3O_2$) and phosphoric acid (H_3PO_4). (Any weak acid is slightly ionized.)

From the four examples cited we see that a double displacement reaction will occur if at least one of the following classes of substances is formed by the reaction:

1. A precipitate.
2. A gas.
3. A slightly ionized compound (a weak electrolyte).
4. A non-ionized compound (a nonelectrolyte).

The following solubility rules are helpful in predicting precipitates that might form when solutions are mixed.

Solubility Rules for Ionic Compounds at 25°C

1. A compound will be soluble if it contains at least one of the following ions: Li^+, Na^+, K^+, NH_4^+, NO_3^-, or $C_2H_3O_2^-$.

2. A compound containing Cl^-, Br^-, or I^- is soluble **unless** the cation is Ag^+, Pb^{2+}, or Hg_2^{2+}.

3. Compounds that contain SO_4^{2-} are soluble **except** those sulfates that also contain Ba^{2+}, Ag^+, Pb^{2+}, or Ca^{2+}.

4. Most other ionic compounds are insoluble (in other words, the amount of solute that dissolves is extremely small and will be regarded as negligible). Compounds containing S^{2-}, CO_3^{2-}, PO_4^{3-}, and OH^- are insoluble unless the cation is one of those listed in #1 above.

Safety Precautions:

- Wear safety goggles.
- Silver nitrate solutions will stain skin and clothing. If you notice any silver nitrate splashing your skin or clothing, rinse it off immediately. Silver nitrate is a colorless solution that is light-sensitive. The stains don't show up immediately, but you will see them by the next day! Stains on your skin will eventually wear off as your skin wears off; stains on your clothing will be permanent. It is therefore best to wear old clothing for this lab.

Waste Disposal:

- While you are doing the experiment, pour your waste into a beaker. When you are finished with the experiment, pour the contents of the waste beaker into the **INORGANIC WASTE** container (with a blue label) in the fume hood.

Procedure

Each part of the experiment (except number 12) consists of mixing equal volumes of two solutions in a test tube. Make sure the test tubes are clean and have been rinsed with deionized water and shaken dry. **Use about one milliliter of each solution**. It is not necessary to measure each volume accurately (why?). Record your observations at the time of mixing. Where there is no visible evidence of reaction, feel each tube, or check with a thermometer, to determine if heat is evolved. In each case where a reaction has occurred, complete and balance the overall equation, properly indicating precipitates and gases, and write the net ionic equation for the reaction. When there is no evidence of reaction, write the words "no reaction" as the right-hand side of the equation.

1. Mix 0.1 M sodium chloride ($NaCl$) and 0.1 M potassium nitrate (KNO_3) solutions.
2. Mix 0.1 M sodium chloride ($NaCl$) and 0.1 M silver nitrate ($AgNO_3$) solutions.
3. Mix 0.1 M sodium carbonate (Na_2CO_3) and 6 M hydrochloric acid (HCl) solutions.
4. Mix 10% sodium hydroxide ($NaOH$) and 6 M hydrochloric acid (HCl) solutions.
5. Mix 0.1 M barium chloride ($BaCl_2$) and 3 M sulfuric acid (H_2SO_4) solutions.
6. Mix 6 M ammonia (NH_3) and 3 M sulfuric acid (H_2SO_4) solutions **in the hood**.
7. Mix 0.1 M copper (II) sulfate ($CuSO_4$) and 0.1 M zinc nitrate ($Zn(NO_3)_2$) solutions.
8. Mix 0.1 M sodium carbonate (Na_2CO_3) and 0.1 M calcium chloride ($CaCl_2$) solutions.
9. Mix 0.1 M copper (II) sulfate ($CuSO_4$) and 0.1 M ammonium chloride (NH_4Cl) solutions.
10. Mix 10% sodium hydroxide ($NaOH$) and 6 M nitric acid (HNO_3) solutions.
11. Mix 0.1 M iron (III) chloride ($FeCl_3$) and 10% sodium hydroxide ($NaOH$) solutions.
12. Add 1 g of solid sodium bicarbonate ($NaHCO_3$) to 3 mL of water and shake to dissolve. Add about 1 mL of 1 M hydrochloric acid (HCl) solution, dropwise, using a medicine dropper.

<u>Questions</u>

1. The formation of what four classes of substances caused double displacement reactions to occur in this experiment?
2. Write the equation for the decomposition of sulfurous acid (H_2SO_3).
3. Using four criteria for double displacement reactions, together with the solubility table, predict whether a double displacement reaction will occur in each example below. If reaction will occur, complete and balance the overall equation, properly indicating gases and precipitates. Then write the net ionic equation for the

reaction. If you believe no reaction will occur, write “no reaction” as the right-hand side of the equation. All reactants are in aqueous solution.

a. K_2S + $CuSO_4$ $\rightarrow$
b. KOH + NH_4Cl $\rightarrow$
c. $(NH_4)_2SO_4$ + NaCl $\rightarrow$
d. $CoCl_3$ + NaOH $\rightarrow$
e. Na_2CO_3 + HNO_3 $\rightarrow$

4. Using the solubility rules, predict the products and write the balanced overall equations (including phase symbols) for the following reactions. Write the net ionic equations.

a. Mix aqueous cobalt (III) nitrate with aqueous sodium sulfide.
b. Mix aqueous phosphoric acid with aqueous potassium hydroxide.
c. Mix aqueous ammonium phosphate with aqueous calcium chloride.
d. Mix aqueous potassium bicarbonate with aqueous hydrochloric acid.