Double Replacement Reactions and Solubility
Net Ionic Equations

Introduction

Precipitation reactions, a type of double replacement reaction, are widely used to prepare new compounds and analyze their purity. Precipitation reactions occur when aqueous solutions of ionic compounds are combined and a new ionic compound, which is insoluble in water, is produced. The result is the formation of a precipitate, a solid which settles out of the solution. By carrying out a series of possible double replacement reactions and observing which combinations produce precipitates, we should be able to determine some general rules of solubility for ionic compounds in water.

Concepts

- Double replacement reactions
- Net ionic equations
- Molecular equations
- Solubility rules

Background

Double replacement reactions occur when two ionic compounds (or compounds that break apart to form ions in aqueous solution) exchange ions to form new compounds. Double replacement reactions generally occur in one direction only in response to a "driving force," which provides a reason for the reaction to occur. Two important driving forces for reactions between ions in aqueous solution are formation of a solid (precipitation reactions) and formation of a gas or stable molecular product (acid–base neutralization reactions). In the absence of a driving force, ionic compounds will remain dissolved in solution and no chemical reaction will take place among the dissolved ions.

When an ionic compound dissolves in water, the crystalline solid dissociates or separates into its corresponding cations and anions. For example, silver nitrate dissociates into silver cations and nitrate anions (Equation 1).

$$AgNO_3(s) \rightarrow Ag^+(aq) + NO_3^-(aq)$$ *Equation 1*

When solutions containing two ionic compounds are mixed, two new combinations of cations and anions are possible. In some cases, the cation from one compound and the anion from the other compound may combine to form an insoluble product, which is called a precipitate. For example, when sodium bromide is added to a solution of silver nitrate, there are four different ions present and two new combinations of cations and anions—sodium nitrate and silver bromide—are possible. Sodium nitrate is soluble in water and its ions remain dissolved in solution. Silver bromide is insoluble in water and precipitates from solution as a solid (Equation 2).

$$AgNO_3(aq) + NaBr(aq) \rightarrow AgBr(s) + NaNO_3(aq)$$ *Equation 2*

Equation 2 summarizes the double replacement reaction between silver nitrate and sodium bromide. It is called a *molecular equation* because the reactants and products are represented by their molecular formulas. The actual reaction process can be better understood if we rewrite the equation to represent the ionic compounds as they exist in solution, that is, in the form of their dissolved ions.

The *complete ionic equation* for the double replacement reaction of silver nitrate and sodium bromide is shown in Equation 3. Because silver bromide is not soluble, it is best represented by its formula, AgBr.

$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Br^-(aq) \rightarrow AgBr(s) + Na^+(aq) + NO_3^-(aq) \qquad \textit{Equation 3}$$

Notice that some ions (Na^+ and NO_3^-) appear on both sides of the equation in Equation 3. These ions are referred to as *spectator ions* because they do not participate in the overall reaction. Just as in algebra, where a term that appears on both sides of an equation may be "cancelled out," we can do the same thing with the complete ionic equation shown in Equation 3. Subtracting the spectator ions from both sides gives the *net ionic equation* for the precipitation reaction of silver bromide (Equation 4).

$$Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s) \qquad \textit{Equation 4}$$

Double replacement reactions may be used to determine some general *solubility rules* for ionic compounds in water. If a precipitation reaction occurs when solutions of ionic compounds are mixed, then at least one of the two possible products formed by the "ion exchange" reaction must be insoluble in water. In this experiment, various ionic solutions will be mixed two at a time to determine which combinations will form precipitates. Knowing the ions that are present will make it possible to predict the solubility pattern of different ionic compounds.

Experiment Overview

The purpose of this experiment is to carry out a series of possible double replacement reactions and analyze the results to formulate some general rules of solubility for ionic compounds. The solutions that will be mixed fall into two categories. The "anion testing solutions" contain sodium salts with six different anions (sodium carbonate, sodium chloride, sodium hydroxide, etc.), while the "cation testing solutions" contain nitrate salts with eight different cations (aluminum nitrate, barium nitrate, calcium nitrate, etc.) All sodium salts and all nitrate salts are soluble in water—therefore, any precipitate that is observed when an anion and cation testing solution are combined should be easy to identify.

Pre-Lab Questions

1. Read the entire *Procedure,* including the *Safety Precautions*. Which solutions used in this experiment are described as skin and eye irritants? What does this mean? What precautions are used to protect against these hazards?

2. Solutions of calcium, zinc, and lead nitrate were mixed pairwise with sodium iodide, sodium sulfate, and sodium chromate using the procedure described in this experiment (see the following table of results). Write a molecular equation and a net ionic equation for each double replacement reaction that produced a precipitate.

	Calcium Nitrate, $Ca(NO_3)_2$	**Lead Nitrate, $Pb(NO_3)_2$**	**Zinc Nitrate, $Zn(NO_3)_2$**
Sodium Chromate, Na_2CrO_4	PPT	PPT	PPT
Sodium Iodide, NaI	NR	PPT	NR
Sodium Sulfide, Na_2S	NR	PPT	PPT

3. What are the advantages and disadvantages of using molecular equations, complete ionic equations, and net ionic equations to describe double replacement reactions?

Materials

Anion Testing Solutions, 0.1 M

Sodium carbonate, Na_2CO_3, 3 mL
Sodium chloride, NaCl, 3 mL
Sodium hydroxide, NaOH, 3 mL
Sodium iodide, NaI, 3 mL
Sodium phosphate, Na_3PO_4, 3 mL
Sodium sulfate, Na_2SO_4, 3 mL

Cation Testing Solutions, 0.1 M

Aluminum nitrate, $Al(NO_3)_3$, 2 mL
Ammonium nitrate. NH_4NO_3, 2 mL
Barium nitrate, $Ba(NO_3)_2$, 2 mL
Calcium nitrate, $Ca(NO_3)_2$, 2 mL
Copper(II) nitrate, $Cu(NO_3)_2$, 2 mL
Iron(III) nitrate, $Fe(NO_3)_3$, 2 mL
Silver nitrate, $AgNO_3$, 2 mL
Zinc nitrate, $Zn(NO_3)_2$, 2 mL

Equipment

Cotton swabs
Distilled water and wash bottle
Labels, adhesive, 14
Paper towels
Pipets, thin stem, 14
Pipet holder (24-well plate)
Reaction plate, 96-well
Toothpicks

Safety Precautions

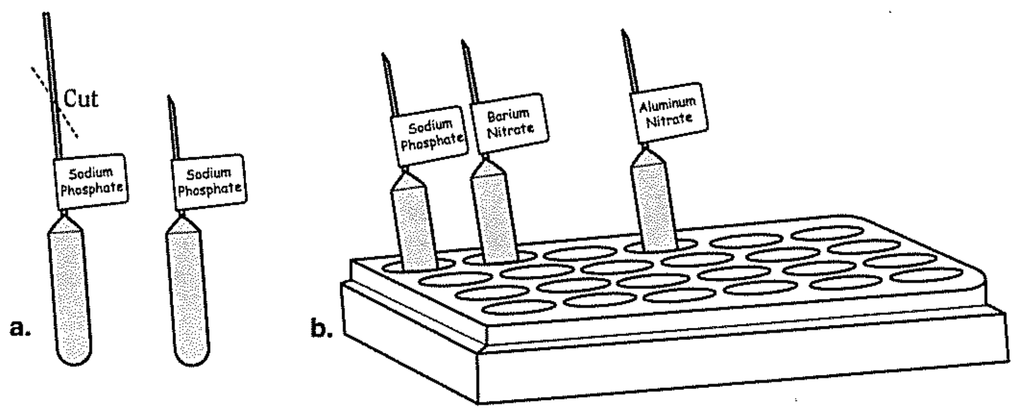
Ammonium nitrate, barium nitrate, copper(II) nitrate, and silver nitrate solutions are slightly toxic by ingestion. Silver nitrate, sodium carbonate, and sodium hydroxide solutions are skin and eye irritants; silver nitrate will also stain skin and clothing. Avoid contact of all chemicals with eyes and skin. Wear chemical splash goggles and chemical-resistant gloves and apron. Wash hands thoroughly with soap and water before leaving the lab.

Procedure

Preparation

Steps 1–4 may only need to be done once by the first class section doing the experiment.

1. Obtain 14 thin stem pipets and 14 blank labels. Using a permanent marker or wax pencil, write the names of the 14 solutions listed in the *Materials* section on the labels. Write on only one-half of each label—the labels will be folded in half around the pipets.
2. Fold the labels in half around the pipet stems just above the bulbs, as shown in Figure 1a.
3. Cut the stem of each pipet at a 45° angle about 5 cm from the bulb (Figure 1a). The shorter stems will make it easier to deliver uniform-size drops.
4. Fill each pipet with the appropriate solution and store the filled pipets stem-side-up in the 24-well plate, as shown in Figure 1b.
5. Place a clean 96-well reaction plate on top of a sheet of black paper as shown below. Each well is identified by a unique combination of a letter and a number—horizontal rows are identified by the letters A–H and vertical columns are numbered from 1 to 12 (Figure 2).



Figures 1a and 1b.

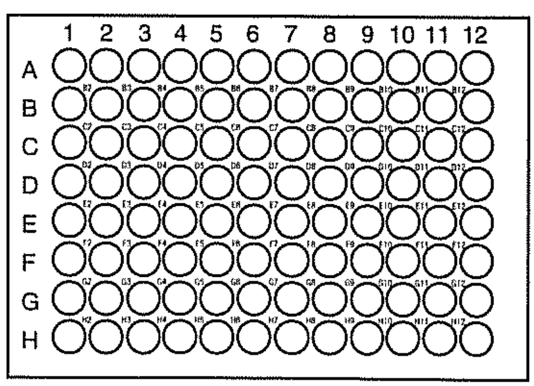


Figure 2. Lettering and Numbering in a 96-Well Plate.

Precipitation Reactions

6. Place 4–5 drops of aluminum nitrate solution, the first "cation testing solution," $Al(NO_3)_3$, into well A1. Hold the pipet vertically to ensure uniform drops. Continue adding 4–5 drops of aluminum nitrate solution to the next five wells in column 1 (wells B1–F1).

7. Using the data table as a guide, place 4–5 drops of the appropriate "cation testing solution" into the first six wells in each vertical column 2–8 (ammonium nitrate into wells A2–F2, barium nitrate into wells A3–F3, etc.). *Note:* Consult the data table frequently and carefully read each label to avoid filling wells with the wrong solution.

8. Add 4–5 drops of sodium carbonate solution, the first "anion testing solution," Na_2CO_3, into well A1.

9. If a precipitate forms or the mixture appears cloudy, write PPT in the correct circle in the data table. If no precipitate forms, write NR (no reaction) in the circle. Remember that not all pairs of solutions will react, and that sometimes the precipitate may take a minute to develop. Use a *clean* toothpick to stir the mixture if the result is not obvious.

10. Continue adding sodium carbonate solution to each filled well in row A (wells A2–A8). When all of the mixtures have been made, go back and record any evidence of reaction in each well as PPT or NR, as described in step 9.

11. In the same manner, add the appropriate "anion testing solution" to each filled well in its row, as shown in the data table (sodium chloride to wells B1–B8, sodium hydroxide to wells C1–C8, etc.) *Note:* Consult the data table frequently and carefully read each label to avoid filling wells with the wrong solution.

12. Record any evidence of reaction in each well as PPT or NR. If there is any doubt about the observations in any well, repeat that test in an empty well on the reaction plate.

Disposal

13. Some of the transition metal and other heavy metal salts used in this experiment should not be discarded into the water supply. Dump the contents of the 96-well reaction plate onto folded paper towels and discard the used paper towels in the wastebasket.

14. Use cotton swabs to clean out any remaining residues in the reaction plate. Thoroughly rinse the reaction plate several times with distilled water.

Name: ______________________________

Class/Lab Period: ______________________________

Double Replacement Reactions

Data Table

	$Al(NO_3)_3$	NH_4NO_3	$Ba(NO_3)_2$	$Ca(NO_3)_2$	$Cu(NO_3)_2$	$Fe(NO_3)_3$	$AgNO_3$	$Zn(NO_3)_2$
Na_2CO_3	A1	A2	A3	A4	A5	A6	A7	A8
NaCl	B1	B2	B3	B4	B5	B6	B7	B8
NaOH	C1	C2	C3	C4	C5	C6	C7	C8
NaI	D1	D2	D3	D4	D5	D6	D7	D8
Na_3PO_4	E1	E2	E3	E4	E5	E6	E7	E8
Na_2SO_4	F1	F2	F3	F4	F5	F6	F7	F8

Post-Lab Questions *(Use a separate sheet of paper to answer the following questions.)*

1. For each combination of an anion and cation testing solution that produced a precipitate, write the name and formula of the solid product. *Hint:* Recall that all sodium salts and all nitrate salts are soluble. Remember to balance the positive and negative charges in the formulas of the products.
2. Write both a balanced molecular equation and a net ionic equation for each precipitation reaction observed in this experiment.
3. (a) Which anion testing solutions produced the fewest (<3) precipitates?

   (b) Salts of these anions may be described as soluble with only a few exceptions. Name the cations that are the *exceptions* to the rule for each anion identified in 3a.
4. (a) Which anion testing solutions produced the most (>5) precipitates?

   (b) Salts of these anions may be described as insoluble with only a few exceptions. Name the cations that are *exceptions* to the rule for each anion identified in 4a.
5. Although you did not test the bromide (Br^-) anion, would you expect most bromide salts ($AlBr_3$, NH_4Br, $BaBr_2$, etc.) to be soluble or insoluble? What is the likely exception to this rule? *Hint:* Use the periodic table to see which of the anions you tested would behave like bromide.
6. Complete the following table to summarize the general solubility rules of ionic compounds. Two entries have been filled in for you.

Ionic Compounds	Soluble or Insoluble	Exceptions
Carbonates		
Chlorides, Bromides, and Iodides		
Hydroxides		
Nitrates	Soluble	None
Phosphates		
Sulfates		
Alkali Metal Salts (Na^+, K^+, etc.)	Soluble	None
Ammonium Salts		