Purpose

To investigate soluble and insoluble ionic compounds and to make predictions regarding solubility.

Learning Objectives

Apply the solubility rules to predict the solubility of ionic compounds.

Carry out and record observations of precipitation reactions.

Write balanced chemical equations for observed double displacement reactions.

Predict double displacement reactions using solubility guidelines.

Write ionic equations for double displacement reactions.

Identify spectator ions in double displacement reactions.

Write net ionic equations for double displacement reactions.

Laboratory Skills

Make and record observations.

Carry out a series of different reactions using a spot plate.

Equipment

- 24-well spot plates
- Dropper bottles

Chemicals

- Sodium carbonate, anhydrous
- Sodium chloride
- Sodium hydroxide
- Sodium nitrate
- Sodium phosphate, dodecahydrate
- Sodium sulfate, anhydrous
- Potassium nitrate
- Calcium nitrate, tetrahydrate
- Barium nitrate
- Aluminum nitrate, nonahydrate
- Silver nitrate
- Copper(II) nitrate, trihydrate
- Iron(III) nitrate
- Lead(II) nitrate

Introduction

Solubility

When a solute is able to dissolve within a solvent, the solute can be described as "soluble" within that solvent. Solubility plays an important role within chemistry, as it is a significant method for adding or removing chemicals from a system. Understanding solubility allows hospital pharmacists to correctly determine if a particular medication will dissolve in an IV solution, or help an environmental chemist remove certain chemicals from water by precipitation. In this lab, students explore a bit about solutions that contain ions that, while individually soluble, may combine to form insoluble substances.

Double Displacement Reactions

Double displacement (or double replacement or metathesis) reactions are common chemical reactions in which two ionic compounds react to form two new compounds. The initial two ionic compounds are soluble in water. When mixed, the cation of each reactant combines with the anion of the other reactant as shown in Reaction IS.1.

$$\mathrm{AX + BY \longrightarrow AY + BX} \qquad \text{(Reaction IS.1)}$$

During a double displacement reaction, there is no change in the charges of individual ions. However, the subscripts for each ion may change because the new compounds must contain the correct ratio of cations and anions to form a neutral compound (total charge of zero). One driving force of a double displacement reactions is the *formation of an insoluble product*, or *precipitate*. This is known as a **precipitation** reaction. Two soluble ionic compounds combine in solution to exchange ions in such a way that an *insoluble* ionic compound forms and precipitates out of solution. In Reaction IS.2, aqueous barium chloride reacts with aqueous sodium sulfate to form solid barium sulfate (the precipitate) and aqueous sodium chloride.

$$\mathrm{BaCl_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + 2\,NaCl(aq)} \qquad \text{(Reaction IS.2)}$$

Another type of double displacement reaction is the reaction between an acid and a base, in which the driving force is the *formation of water*. Thus, there is no precipitate to observe. A strong acid is a soluble ionic compound with a hydrogen cation and a strong base is a soluble ionic compound with a hydroxide anion. When the double replacement occurs, the hydrogen and hydroxide ions combine to form water molecules. The remaining cation and anion combine to produce another ionic compound. In Reaction IS.3, aqueous hydrochloric acid reacts with aqueous sodium hydroxide to form liquid water and aqueous sodium chloride.

$$\mathrm{HCl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaCl(aq)} \qquad \text{(Reaction IS.3)}$$

Double displacement reactions can also result in the *formation of a gaseous product*, as seen in the reaction of aqueous sodium sulfide and hydrochloric acid in Reaction IS.4.

$$\mathrm{2\,HCl(aq) + Na_2S(aq) \longrightarrow H_2S(g) + 2\,NaCl(aq)} \qquad \text{(Reaction IS.4)}$$

The key for distinguishing among the types of double displacement reactions is the *evidence of a reaction*, which can be either the formation of a precipitate, the formation of water (or other small, stable molecule), or the formation of a gas. Of course, it is entirely possible to combine two solutions of ions and observe no sign of a reaction. In these cases, the reactant ions simply exist in solution together, forming a sort of "ion soup" with no formation of a precipitate, water molecules, or a gas. An example of this seen in Reaction IS.5, where no reaction has occurred because the products are all soluble.

$$\mathrm{NaCl(aq) + KNO_3(aq) \longrightarrow KCl(aq) + NaNO_3(aq)} \qquad \text{(Reaction IS.5)}$$

Solubility of Ionic Compounds

The solubility rules in Table IS.1 describe some general patterns of solubility of ionic compound as observed by chemists. These rules or guidelines can be used to predict the solubility of an ionic compound in water. Terms such as "nitrates" and "sulfates" refer to all ionic compounds that have nitrate ions (NO_3^-) and sulfate ions (SO_4^{2-}), respectively as their anions.

Table IS.1: Rules for Predicting the Solubility of Ionic Compounds

Rule	Applies to	Statement	Exceptions
1	Group 1 and ammonium ions	All compounds are soluble.	–
2	Acetates, nitrates	All compounds are soluble.	–
3	Halides (Cl^-, Br^-, I^-)	Most halides are soluble.	Silver, mercury, and lead halides
4	Sulfates	Most sulfates are soluble.	Calcium, strontium, barium, silver, mercury, and lead sulfates
5	Carbonates	Most carbonates are insoluble.	Group IA and ammonium carbonates
6	Phosphates	Most phosphates are insoluble.	Group IA and ammonium phosphates
7	Sulfides	Most sulfides are insoluble.	Group IA and ammonium sulfides
8	Hydroxides	Most hydroxides are insoluble.	Group IA and ammonium hydroxides

Example IS.1

Use the solubility rules to determine if any of the following compounds is likely to be soluble in water: sodium phosphate, calcium carbonate, and lead(II) chloride.

Start by writing the chemical formulas for each compound and applying the solubility rules.

- Sodium phosphate, Na_3PO_4, contains sodium ions and phosphate ions. According to rule 6, most phosphates are insoluble. However, sodium is a group 1 metal and rule 1 states that all ionic compounds containing group 1 metal ions are soluble.

Thus, sodium phosphate is soluble in water.

- Calcium carbonate, $CaCO_3$, contains calcium ions and carbonate ions. According to rule 5, most carbonates are insoluble. There is no additional rule or exception for calcium (group 2) ions.

Thus, calcium carbonate is insoluble in water.

- Lead(II) chloride, $PbCl_2$, contains lead(II) ions and chloride ions. Chloride ions are halides and rule 3 states that most halides are soluble. However, lead is listed as an exception to rule 3.

Thus lead(II) chloride is insoluble in water.

The solubility guidelines are also used to predict if a precipitate will form if two solutions of different ionic compounds are mixed. Recall that precipitation reactions are double displacement reactions and follow the general format shown in Reaction IS.1.

$$AX + BY \longrightarrow AY + BX \qquad \text{(Reaction IS.1, revisited)}$$

The formulas of the compounds in the two solutions are used to predict formulas of possible products, if a reaction were to occur. If any of the possible products is insoluble in water (according to the solubility rules), then a precipitate will form.

Example IS.2

Will a precipitate form if aqueous solutions of lead(II) nitrate and sodium bromide are mixed?

Start by writing a hypothetical equation, to identify the two possible products. A word equation is sufficient at this point. Simply interchange the cations and anions in each compound.

$$\text{lead(II) nitrate} + \text{sodium bromide} \longrightarrow \text{lead(II) bromide} + \text{sodium nitrate}$$

Now apply the solubility rules to determine if either of the possible products will form a precipitate.

Lead(II) bromide: Rule 3 states that most halides are soluble, but lead halides are exceptions. Thus, lead bromide is insoluble in water and will form a precipitate.

Sodium nitrate: Rule 1 states that all compounds containing Group 1 ions are soluble and Rule 2 states that all nitrates are soluble.
Thus, sodium nitrate is soluble and will not form a precipitate.

Writing Ionic Equations

Aqueous Solutions of Ionic Compounds

There are several types of chemical equations used to describe what is happening in a reaction. An **overall equation** gives the formulas of the reactants and products and shows the stoichiometry using coefficients. The overall reaction of barium chloride and sodium sulfate was given in Reaction IS.2:

$$BaCl_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + 2\,NaCl(aq) \qquad \text{(Reaction IS.2, revisited)}$$

Recall that the (aq) after the formula of a compound indicates an aqueous solution of that compound. Ionic compounds, such as $BaCl_2$ and Na_2SO_4 dissociate in water to form hydrated ions, as shown in Reaction IS.6 and Reaction IS.7:

$$BaCl_2(s) \rightleftharpoons Ba^{2+}(aq) + 2\,Cl^-(aq) \qquad \text{(Reaction IS.6)}$$

$$Na_2SO_4(s) \rightleftharpoons 2\,Na^+(aq) + SO_4{}^{2-}(aq) \qquad \text{(Reaction IS.7)}$$

The hydrated ions float around in solution and, when they bump into the right ions, can form new, insoluble ionic compounds.

Ionic Equations

When aqueous barium chloride is mixed with aqueous sodium sulfate, the separate, hydrated ions (See Reaction IS.6 and Reaction IS.7) form the precipitate, barium sulfate. Reactions of separate, hydrated ions can be shown in an **ionic equations**, such as the one showing the formation of barium sulfate, Reaction IS.8.

$$Ba^2 + (aq) + 2\,Cl^-(aq) + 2\,Na^+(aq) + SO_4{}^{2-} \longrightarrow BaSO_4(s) + 2\,Na^+(aq) + 2\,Cl^-(aq) \qquad \text{(Reaction IS.8)}$$

Note that in the ionic equation, the *subscripts* from the formulas in the overall equation become *coefficients* of the separated ions. Not also that any reactants or products listed as solids or liquids remain as compounds. Barium chloride, sodium sulfate, and sodium chloride are soluble in water and are shown as separate, hydrated ions. In Reaction IS.8, the two chloride ions present in $BaCl_2$ become $2Cl^-$. However, barium sulfate is written as a compound because it is insoluble and, thus, does not dissociate to form aqueous ions.

Spectator Ions

Some of the ions in Reaction IS.8 do not change during the reaction, but instead remain as aqueous ions. These ions, which do not participate in the reaction. are called spectator ions. Sodium and chloride ions are the spectators in this reaction (Figure IS.1).

$$Ba^{2+}(aq) + \boxed{2Cl^{-}(aq)} + \boxed{2Na^{+}(aq)} + SO_4^{2-} \longrightarrow BaSO_4(s) + \boxed{2Na^{+}(aq)} + \boxed{2Cl^{-}(aq)}$$

Figure IS.1: Ionic Equation Showing Spectator Ions.

Net Ionic Equations

A **net ionic equation** focuses on just those ions that undergo a change, omitting the spectator ions, as shown in Reaction IS.9.

$$Ba^{2+}(aq) + SO_4{}^{2-}(aq) \longrightarrow BaSO_4(s) \qquad \text{(Reaction IS.9)}$$

This equation shows barium and sulfate ions combining to produce the precipitate, barium sulfate. Ionic and net ionic equations (and all chemical equations) must be balanced for mass and charge. That is, the same number of each type of ion must be present on each side of the arrow.

Example IS.3

Write a net ionic equation for the reaction that occurs when solutions of calcium nitrate and potassium hydroxide are combined, forming solid calcium hydroxide and an aqueous solution of potassium nitrate.

The process of writing a net ionic equation starts with a balanced overall equation. The same number of each type of ion must be present on both sides of the arrow to balance the equations for both charge and mass.
Start by writing correct formulas for the reactants and products:

$$Ca(NO_3)_2(aq) + KOH(aq) \longrightarrow Ca(OH)_2(s) + KNO_3(aq)$$

Add coefficients to balance the overall equation:

$$Ca(NO_3)_2(aq) + 2\,KOH(aq) \longrightarrow Ca(OH)_2(s) + 2\,KNO_3(aq)$$

Now separate the ions in the soluble compounds to write a balanced ionic equation. Recall that the subscripts in the formulas become coefficients for the separated ions and that the coefficients in the overall equation apply to both ions in the formula!

$$Ca^{2+}(aq) + 2\,NO_3{}^{-}(aq) + 2\,K^{+}(aq) + 2\,OH^{-}(aq) \longrightarrow Ca(OH)_2(s) + 2\,K^{+}(aq) + 2\,NO_3{}^{-}(aq)$$

Cross out the spectator ions to write a balanced net ionic equation.

$$Ca^{2+}(aq) + 2\,OH^{-}(aq) \longrightarrow Ca(OH)_2(s)$$

Aqueous calcium ions react with aqueous hydroxide ions to form solid calcium hydroxide.

Procedure

All the solutions used in this experiment are 0.1 M concentrations. Each will be in a dropper bottle. The chemicals include:

- Sodium salts of the following anions: CO_3^{2-}, Cl^-, OH^-, NO_3^-, PO_4^{3-}, and SO_4^{2-}
- Nitrate salts of these cations: K^+, Ca^{2+}, Ba^{2+}, Al^{3+}, Ag^+, Cu^{2+}, Fe^{3+} and Pb^{2+}

1. Obtain a spot plate and make sure it is clean. Contaminants left from previous use can cause false reactions in this procedure. If necessary, use soap and water to wash it first, rinsing three times with RO or DI water. It is not essential to have it completely dry.

2. Divide the metal ions into two groups of four as you set up your mixtures. A spot plate has six columns and four rows. Thus, the six anions will each be added to one column on both spot plates; while the metal ions will each be added to one row.

3. Your first spot plate will contain all six anions:

$$CO_3^{2-}, Cl^-, OH^-, NO_3^-, PO_4^{3-}, \text{and } SO_4^{2-}$$

and these four cations:

$$K^+, Ca^{2+}, Ba^{2+}, Al^{3+}$$

4. Place your spot plate on notebook paper and label the paper for each row and column similarly to the way your report sheet is set up.

5. In each well of the spot plate you will place 2 drops of each of the two specified solutions. Observe and note any signs of a reaction.

6. Record the results on your report sheet. Use **ppt** to indicate the formation of a precipitate. Describe the color and appearance of the precipitate, such as "yellow, chunky ppt" or "white, fluffy ppt." If no precipitate forms, enter NR for "no reaction."

7. Once you have completed all 24 combinations and recorded your observations, take a photo of your spot plate that includes the labels you made on notebook paper.

8. Clean the spot plate by pouring the contents carefully into a large waste beaker. Rinse the plate twice with RO or DI water and pour the rinses into the waste beaker. Wash the plate with soap and water and rinse thoroughly with RO or DI water.

9. label a second sheet of notepaper with the ions for the second spot plate, which contains all six anions:

$$CO_3^{2-}, Cl^-, OH^-, NO_3^-, PO_4^{3-}, \text{and } SO_4^{2-}$$

and these four cations:

$$Ag^+, Cu^{2+}, Fe^{3+} \text{ and } Pb^{2+}$$

10. Complete another 24 combinations of two chemical solutions and record your observations for each, as you did for the first spot plate.

11. Take a photo of the second spot plate including the notebook paper labels.

12. Clean the spot plate by pouring the contents carefully into a large waste beaker. Rinse the plate twice with RO or DI water and pour the rinses into the waste beaker. Wash the plate with soap and water and rinse thoroughly with RO or DI water.

13. Pour the contents of the waste beaker into the metal ion waste container in the hood.

Name: ______________________________

Section: __________ Date: ______________

Report Sheet:

Soluble and Insoluble Salts

First spot plate results:

	CO_3^{2-}	Cl^-	OH^-	NO_3^-	PO_4^{3-}	SO_4^{2-}
K^+						
Ca^{2+}						
Ba^{2+}						
Al^{3+}						

Second spot plate results:

	CO_3^{2-}	Cl^-	OH^-	NO_3^-	PO_4^{3-}	SO_4^{2-}
Ag^+						
Cu^{2+}						
Fe^{3+}						
Pb^{2+}						

Post-Lab Questions

1. Consider the colored precipitates you produced and the cations in those reactions. Identify the colorful cations and describe their location on the periodic table. What do they have in common? Select the statement that best answers the question:

   A. They are from the same row on the periodic table.
   B. They are from the same column on the periodic table.
   C. They are all nonmetals.
   D. They are transition metals.
   E. They are all alkali metals.

2. If you believe that you have a water sample that is contaminated with barium:

   A. What could you add to it to test it?

   B. What would be the sign that it is contaminated?

3. Based on the solubility observations, which of the following pairs of cation could be distinguished by the addition of sodium chloride to the solutions?

   A. Barium and lead
   B. Barium and aluminum
   C. Lead and sliver
   D. Iron and calcium

4. Based on the solubility observations, would you expect potassium and bromide to form a precipitate?

5. Based on the solubility observations, would you expect sodium and nitrate to form a precipitate?

6. Write balanced chemical equations as indicated for any observed reactions. Write NR if no reaction was observed.

   A. Ba^{2+} and CO_3^{2-}, overall equation ______________________________

   B. Ba^{2+} and CO_3^{2-}, ionic equation ______________________________

   C. Ba^{2+} and CO_3^{2-}, net ionic equation ______________________________

   D. Ba^{2+} and PO_4^{3-}, overall equation ______________________________

   E. Ba^{2+} and Cl^-, ionic equation ______________________________

   F. Ba^{2+} and OH^-, net ionic equation ______________________________