Lab 6: Solubility

1. Modeling Molecular Structure

1.1. Electrolytes & Precipitation Reactions

1.2. Electrolytes verses non-electrolytes.

1.3. Soluble or non-soluble

1.4. Complete ionic equations and net-ionic equations

1.5. Solubility rules

2. Learning Objectives

2.1. Define the differences between electrolytes and non-electrolytes.

2.2. Carry out and record observations of electrical conductivity.

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

2.4. Carry out and record observations of precipitation reactions.

2.5. Write complete ionic equations and net-ionic equations.

2.6. Write balanced chemical equations for observed double displacement reactions.

3. Equipment

3.1. Laminated target sheets

3.2. Dropper pipettes

4. Chemicals

4.1. Sodium carbonate, anhydrous

4.2. Sodium chloride

4.3. Sodium hydroxide

4.4. Sodium nitrate

4.5. Sodium phosphate · dodecahydrate

4.6. Sodium sulfate, anhydrous

4.7. Potassium nitrate

4.8. Calcium nitrate · tetrahydrate

4.9. Barium nitrate

4.10. Aluminum nitrate · nonahydrate

4.11. Silver nitrate

4.12. Copper(II) nitrate · trihydrate

4.13. Iron(II) nitrate ·hexahydrate

4.14. Lead(II) nitrate

5. Additional Resources

6. Introduction

6.1. Electrolytes

Electrolytes are a category. The word *electrolyte* derives from Ancient Greek ήλεκτρο- (*ēlectro-*), prefix related to electricity, and λυτός (*lytos*), meaning "able to be untied or loosened".[1] As the name describes, there is a separation of charges for electrolytes. Specifically, in chemistry we refer to this separation of charges as the dissolving into ions of a molecule whether it be a salt, an acid, or a base. Note, that the separation of charges is the key idea, this says nothing about the physical state of the solution and that the solution can be solid, liquid, gas, or even plasma.

What matters with electrolytes is the separation of ions into cation and anion. This is very important to many processes. In galvanic cells and batteries, the electrolytes are fundamental to the movement of electrical current. Solid-state electrolytes are commonly found in electronics. The lithium-ion battery is a good example of a solid state electrolyte and there is considerable active research being conducted on this topic at our university and many others throughout the world.

In biological systems, electrolytes have incredible physiological importance. All known multicellular lifeforms have a dynamic (always in motion) electrolyte equilibrium. In larger lifeforms, the electrolyte equilibrium plays major roles in inter- and intracellular signaling including muscle and nerve function. In the human body, serious electrolyte disruption, such as dehydration and overhydration, may lead to cardiac and neurological complications and, unless they are rapidly resolved, will result in a medical emergency.

[1] https://en.wikipedia.org/wiki/Electrolyte

For this lab, we will be demonstrating the conductivity of deionized water and common vegetable oil. Deionized water is water that has been mechanically filtered or processed to remove impurities, typically dissolved metal cations. Vegetable oil is an organic, meaning carbon based, solvent. Vegetable oil is non-polar with few bond dipoles and a negligible net dipole over the entire molecule. Without dissolved ions separated out into cations and anions, there should be very little, if any, conductivity.

We will add some salt to both the DI water and the vegetable oil and see if this changes. If it changes, that means that our salt has split into cations and anions. The charge separation is what allows the electrical conduction to occur.

6.2.Solubility

Solutions consist of two or more substances combined to form a homogeneous mixture. The components present in lesser quantities are *solutes* dissolved in the larger quantity component, the *solvent*. The most common solutions involve solid solutes dissolved in water to form aqueous solutions.

When a solute can 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.

6.3.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.

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

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.

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

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.

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

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.

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

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 1 and ammonium carbonates
6	Phosphates	Most phosphates are insoluble.	Group 1 and ammonium phosphates
7	Sulfides	Most sulfides are insoluble.	Group 1 and ammonium sulfides
8	Hydroxides	Most hydroxides are insoluble.	Group 1 and ammonium hydroxides

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.

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

6.4.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.

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, Na3PO4, 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, CaCO3, 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, PbCl2, 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.

7. Procedure

7.1. Precipitation experiment

7.1.1. All the solutions used in this experiment are 0.2 M concentrations. Each will be in a dropper bottle.

7.1.2. Thoroughly clean the laminated data table.

7.1.3. Add **1 drop** (~1 mL) of each reactant to each square as indicated on the table right in the center of the target.

- The black and white colors on the circle target are there to make any precipitate more visible.
- Too many drops will contaminate the other sites, use only one drop of each reactant so that the total volume in each site is ~2 mL.

7.1.4. 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."

7.1.5. Clean the laminated data table by wiping with a paper towel and washing with DI water.

7.1.6. Check that the results are recorded the results on your report sheet and write the net ionic equations as appropriate.

7.1.7. Clean the laminated sheet 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.

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

7.2. Conductivity demonstration.

7.2.1. Observe and record the results on the datasheet under 8.2.

8. Data Sheet
   6.5. Solubility tests
   6.6. Record whether or not there is a precipitate and if there is, write the net ionic equation. Go ahead and use ppt to mean precipitate.

	I^-	Cl^-	OH^-	NO_3^-	SO_4^-	CO_3^{2-}	PO_4^{3-}
Na^+		No PPT					
K^+							
Ag^+							
Mg^{2+}							
Ca^{2+}							
Fe^{2+}							
Cu^{2+}							
Pb^{2+}							

8.1. Electrolytes

8.1.1. Does the deionized (DI) water conduct electricity? Why or why not?

8.1.2. Does the deionized (DI) water conduct electricity when salt is added? Why or why not?

8.1.3. Does the vegetable oil conduct electricity? Why or why not?

8.1.4. Does the vegetable oil conduct electricity when salt is added? Why or why not?

9. Post-Lab Questions

9.1. Pick a reaction that formed a precipitate in 8.1:

9.1.1. Write out the **full ionic equation** with all of the phases ((s) = solid, (l) = liquid, (aq) = aqueous (in water), (g) = gas) noted in the equations:

- Cation separation:

- Anion separation:

- Cation and anions solutions together to form product:

- Identify the spectator ions:

9.1.2. Write out the **net ionic** equation:

9.2. Pick a different reaction that formed a precipitate in 8.1:

9.2.1. Write out the **full ionic equation** with all of the phases ((s) = solid, (l) = liquid, (aq) = aqueous (in water), (g) = gas) noted in the equations:

- Cation separation:

- Anion separation:
- Cation and anions solutions together to form product:
- Identify the spectator ions:

9.2.2. Write out the **net ionic** equation:

9.3. Thought experiment: is it possible to make a non-polar electrolyte? Explain your answer.

10. Summary and Conclusions