Chemical Reactions

Experiment #6

Objective: To observe different types of chemical reactions and how concentration of reactants or temperature can affect the rate of a chemical reaction.

Introduction

There are many ways of classifying chemical reactions, such as combination, decomposition, single displacement, double displacement and combustion. Many metal ions form salts with chloride (Cl^-) ion, sulfate (SO_4^{2-}) ion, nitrate (NO_3^-) ion and other anions. Very often these salts are hydrates, meaning they also have water bound to the metal ions. When the metals are dissolved in water they may form hydrate complex ions, such as $[Cu(H_2O)_5]^{2+}$ or $[Co(H_2O)_6]^{2+}$. It is easy to decompose a hydrate salt, such as $CuSO_4{\cdot}5H_2O$, by heating it up to drive off the water molecules, this would be a dehydration reaction or decomposition of the hydrate salt.

Cu^{2+} ions have water molecules bound to them in aqueous solution. These hydrated ions may react with other ions (e.g., OH^-) or molecules (e.g., NH_3), which may change their solubility or color. The metal ion salts usually form insoluble compounds with the appropriate number of hydroxide (OH^-) ions to balance the charge on the metal ion when a base such as sodium hydroxide is added. When ammonium hydroxide is added, several metal ions form complexes with ammonia (NH_3) bound to the metal ion, which does not neutralize the charge on the metal ion and a charged complex ion is formed. The number of ammonia molecules bound to the metal ion can vary and these ammine complexes are usually soluble in water because of the net charge on the ions. You will observe some of these reactions where ammonia displaces water bound to the metal ion or displaces hydroxide ion from the insoluble hydroxide compound.

Some reactions are classified as oxidation-reduction reactions if there are changes in the oxidation state of some components of the reaction. Transition metal ions such as copper, cobalt, iron and chromium are commonly involved in oxidation-reduction reactions, which occur in biological systems when we metabolize sugars and fats for energy. Combustion reactions are oxidation-reduction reactions, where oxygen is the oxidizing agent. In many biological systems, oxygen is also the oxidizing agent, although there are transition metals and other cofactors involved in the reactions. In this experiment you will study an oxidation-reduction reaction, which is also a simple displacement of zinc metal and copper(II) ions to form zinc(II) ions and copper metal.

Materials: Hot plate; solid copper sulfate pentahydrate; 1.0 M ammonium hydroxide solution; 1.0 M sodium hydroxide solution; 1.0 M hydrochloric acid solution; 0.1 M copper sulfate solution; 0.1 M barium chloride solution; 0.1 M sodium chromate solution; 0.1 M sodium sulfate solution; 0.01 M iron(III) ammonium sulfate solution; 0.01 M ammonium thiocyanate solution.

Procedure

Use disposable plastic Beryl pipets for addition of small volumes (up to 2 mL) of solutions in these experiments. Rinse the pipets by drawing in deionized water and squirting it out in the sink or a waste flask when changing from one solution to another to minimize contamination. Always hold the pipet upright when transferring solutions. Minimize the plastic waste you generate. Discard all test solutions in the hazardous waste bottle in the hood. Do not pour them in the sink!

A. A Decomposition Reaction

1. Weigh a clean, dry 50 mL beaker to the nearest 0.01 g and record its mass on the report sheet.

2. Add about 0.3 to 0.4 g of copper(II) sulfate pentahydrate ($CuSO_4{\cdot}5H_2O$) to the beaker and reweigh it to the nearest 0.01 g and record the mass on the report sheet. Note the color and hue of the copper sulfate crystals on the report sheet.

3. Put the beaker on the hot plate and place a watch glass over the beaker like a dome. Set the temperature of the hot plate to about 180°C and observe whether there is any moisture condensation on the watch glass. After you have seen moisture on the watch glass, remove the watch glass from the beaker and continue heating the crystals until they become mostly white. Record your observations on the report sheet.

4. After most of the crystals have become white or very pale blue, remove the beaker from the hot plate and place it on the bench to cool. When the beaker has cooled, reweigh the beaker with the copper(II) sulfate in it, record the mass on the report sheet. Was there any loss of mass? What substance may have been lost? Where did it go?

5. After you have reweighed the beaker add several drops of deionized water to the beaker and note what happens to the whitish crystals as you add the water. Record your observation on the report sheet and answer the questions for this part.

B.Reactions of Cu^{2+} ions (Cu(II)) with bases (double replacement reactions)

1. Use a plastic pipet to add about 2 or 3 mL of 1.0 M ammonium hydroxide solution (NH_4OH) to a clean **small** test tube (about 1/4 full). Note the odor of the ammonium hydroxide solution but do not get your nose too close to the test tube. You should hold the test tube several inches in front of your face and wave your free hand over the top of the test tube toward your nose until you notice the smell of ammonia. You will use this solution below.

2. Rinse the plastic pipet with water and add about 2 or 3 mL of 1.0 M sodium hydroxide (NaOH) to another clean **small** test tube and notice whether it has any odor like the ammonium hydroxide solution. You can place the plastic pipet in the test tube containing ammonium hydroxide and use it for either the ammonium hydroxide or sodium hydroxide solutions when you add them below.

3. Use a clean plastic pipet and add about 2.0 mL of 0.1 M copper sulfate ($CuSO_4$) solution to two clean **large** test tubes and record the color of the solution on the report sheet.

4. Add drops of the ammonium hydroxide solution to one of the tubes containing copper sulfate. Notice any changes in the solution after adding a few (2 to 4) drops of ammonium hydroxide

solution. You should see a precipitate of copper hydroxide form. Carefully waft the vapors from the large test tube containing copper sulfate and ammonium hydroxide mixture and notice whether you smell any ammonia and record your observations on the report sheet.

5. After recording your observation, continue adding drops of ammonium hydroxide solution until the pale blue precipitate dissolves (shake or stir the tube as you add ammonium hydroxide, it should require less than 1 mL of ammonium hydroxide). Notice whether there is any odor of ammonia coming from the tube containing the copper sulfate-ammonium hydroxide mixture. Record the final observations in terms of color or precipitate in the table for this part on the report sheet.

6. To the same test tube containing copper sulfate-ammonium hydroxide mixture, add about 2 mL of 1.0 M hydrochloric acid (HCl) solution, mix well and notice any changes in the color of the solution or whether there is any precipitate. Record your observations on the report sheet.

7. Add drops of sodium hydroxide solution to the second tube containing the copper sulfate. Notice any changes in the solution after adding a few drops of sodium hydroxide solution. Again, you should see a precipitate of $Cu(OH)_2$ form.

8. Continue adding sodium hydroxide solution until you have added about 1 mL of sodium hydroxide to the copper sulfate and note whether there is any precipitate or color changes after the initial few drops of sodium hydroxide that were added. Record your observations on the report sheet.

9. Using the test tube containing copper sulfate and sodium hydroxide mixture, add about 2 mL of 1.0 M hydrochloric acid (HCl) solution, mix well and notice any changes in the color of the solution or whether there is any precipitate. Record your observations on the report sheet.

C. A single replacement/oxidation-reduction reaction

1. Obtain a paper clip and straighten it the best you can. A paper clip is usually made of steel (iron) with a zinc coating to prevent rust.

2. Add about 2 mL of 0.1 M copper sulfate ($CuSO_4$) solution to a clean **small** test tube (about 1/4 full). Place the paper clip in the test tube and allow it to stand for about 5 to 10 min. You can begin the next part while you are waiting. Notice any change in the wire paper clip that is in contact with the copper sulfate solution compared to the part that is not in the solution. Record your observations on the report sheet and answer the questions for this part.

D. Some double displacement reactions

1. Add about 1 mL of 0.1 M barium chloride ($BaCl_2$) solution to a clean large test tube and then add about 1 mL of 0.1 M sodium chromate (Na_2CrO_4) solution and mix well by holding the top of the test tube with your thumb and one finger of one hand and tap the bottom of the test tube with a finger of the other hand. Record your observations on the report sheet, noting any precipitation or

color changes.

2. Add about 2 mL of 1.0 M hydrochloric acid (HCl) to the test tube and note what happens to the precipitate. Record your observations on the report sheet, noting any color changes, and answer questions for this part.

3. Add about 1 mL of barium chloride solution to another clean large test tube and then add about 1 mL of 0.1 M sodium sulfate (Na_2SO_4) solution and mix as you did for the previous mixture. Record your observations on the report sheet, noting any precipitation or color changes.

4. Add about 2 mL of 1.0 M hydrochloric acid (HCl) to the test tube and note what happens to the precipitate. Record your observations on the report sheet, noting any color changes and answer questions for this part.

5. Add 3 mL of deionized water and 1 mL of 0.01 M iron(III) ammonium sulfate ($Fe(NH_4)(SO_4)_2$) solution to a clean **<u>large</u>** test tube and record the color and appearance (cloudy or clear) of this solution on the report sheet.

6. Add 1 mL of 0.01 M ammonium thiocyanate (NH_4SCN) to the test tube and mix well. Note any color change for this solution and record the ratio of Fe^{3+} ion to SCN^- ion in the table on the report sheet.

7. Add another 1 mL of 0.01 M NH_4SCN solution to the same test tube and record any color change or change in color intensity and the ratio of Fe^{3+} ion to SCN^- ion in the table on the report sheet.

8. Add an additional 2 mL of 0.01 M NH_4SCN solution to the same test tube and record any color change or change in color intensity and the ratio of Fe^{3+} ion to SCN^- ion in the table on the report sheet. This should make a total of 4 mL of NH_4SCN solution added to 1 mL of $FeCl_3$ solution. Record your observations on the report sheet and answer the questions for this part.

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Chemical Reactions
Experiment #6 **Prelab Exercise**

1. Write a balanced chemical equation for the oxidation of zinc metal (Zn) by copper(II) ions (Cu^{2+}) in a copper sulfate solution. Note that copper(II) sulfate completely dissociates in aqueous solution and forms Cu^{2+} (aq) and SO_4^{2-} (aq) in water. Zinc metal is a solid and should be represented as Zn (s).

2. Write a chemical equation for the decomposition of limestone ($CaCO_3$) by heating to high temperature to drive off carbon dioxide. The other product of this decomposition reaction is called lime, a component of cement.

3. Write a chemical equation for the combustion of methane (natural gas) when a gas stove or bunsen burner is lit.

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4. Draw a structural formula with Cu^{2+} ion bound to 5 water molecules. Would you expect the Cu^{2+} ion to be bound to the oxygen or hydrogen of water?

5. Draw a structural formula with Cu^{2+} ion bound to 4 ammonia (NH_3) molecules. Would you expect the Cu^{2+} ion to be bound to the nitrogen or hydrogen of ammonia?

Name ____ Section ______________

Observing Different Types of Chemical Reactions
Experiment #6 **Data & Report Sheet**

A. A decomposition reaction

Mass of empty beaker _______________ g

Mass of beaker with $CuSO_4 \cdot 5\ H_2O$ before heating ______________________ g

Mass of $CuSO_4 \cdot 5\ H_2O$ in the beaker ______________________ g

Mass of beaker with copper(II) sulfate after heating ______________________g

Mass of water lost upon heating the $CuSO_4 \cdot 5\ H_2O$ _______________________g

What is the percent of mass lost from the total mass of $CuSO_4 \cdot 5\ H_2O$? __________________

A-1. What is the formula weight for $CuSO_4 \cdot 5\ H_2O$? _________________

A-2. What is the formula weight for $CuSO_4$? _________________

A-3. What percent of $CuSO_4 \cdot 5\ H_2O$ is water? ________________

A-4. How does the percent loss you observed compare with the theoretical percent of water in $CuSO_4 \cdot 5\ H_2O$? Explain any differences in your observation compared to the theoretical value.

A-5. What is the color/hue of copper(II) sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$)?

A-6. What is the color of the crystals after heating?

A-7. Write a chemical equation for the decomposition reaction occurring in the beaker as you heat copper(II) sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$).

A-8. Write a chemical equation for the hydration reaction occurring in the beaker when water is added to the dehydrated crystals? Is it the reverse reaction for step 4 above?

Name ____ Section ______________

B. Some reactions of transition metal ions (double replacement reactions)

Describe the color and solubility (clear solution or cloudy or formation of precipitate) when ammonium hydroxide is added to $CuSO_4$ solution and when sodium hydroxide is added to $CuSO_4$ solution.

	With Ammonium Hydroxide	With Sodium Hydroxide
Original color of the solution		
Color and solubility after a few drops of hydroxide solution		
Final color and solubility after excess hydroxide solution		
Final color and solubility after 2 mL of hydrochloric acid (1 M HCl) is added.		

Cu^{2+} ions have water molecules bound to them in aqueous solution. These hydrated ions may react with other ions (e.g., OH^-) or molecules (e.g., NH_3), which may change their solubility or color.

B-1. How many water molecules may be bound to the Cu^{2+} ion in water? _____________

B-2. Show the replacement of 2 out of 5 water molecules bound to copper(II) ions in aqueous solution replaced by 2 hydroxide ions (OH^-). Write the formula for copper(II) hydroxide that forms a neutral insoluble precipitate in water.

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B-2. Write the formula for a copper(II) tetraammine complex that has 4 ammonia (NH_3) molecules attached to the Cu^{2+} ion. Indicate any charge that may be present on the complex ion.

B-3. Are the sulfate ions, ammonium ions or sodium ions involved in either of these replacement reactions when ammonium hydroxide or sodium hydroxide is added to $CuSO_4$ solution? If so, indicate which ones and how.

C. An oxidation-reduction reaction

C-1. Describe what happens to the paper clip when it is immersed in the copper(II) sulfate solution.

C-2. Write a chemical equation for the reaction of copper(II) ions reacting with the zinc metal coating on the paper clip.

C-3. What substance is oxidized in this reaction?

C-4. What substance is reduced in this reaction?

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D. More double replacement reactions

Fill in the table with your observations regarding the solutions that are mixed together in this part. A change in color or a change in solubility (formation of a precipitate) would indicate that a chemical reaction has taken place.

	Original solution color	Mixture	
		Color	Solubility
$BaCl_2$			
Na_2CrO_4			
+ HCl			
$BaCl_2$			
Na_2SO_4			
+ HCl			

D-1. Write a chemical equation for the reaction of $BaCl_2$ with Na_2CrO_4 indicating what ions are in aqueous solution (aq) and what compounds are insoluble solids (s).

D-2. Write a chemical equation for the reaction of $BaCl_2$ with Na_2SO_4 indicating what ions are in aqueous solution (aq) and what compounds are insoluble solids (s).

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D-5. Describe the color and appearance of the iron(III) ammonium sulfate ($Fe(NH_4)(SO_4)_2$) solution prior to adding any ammonium thiocyanate in the table below for 0 mL NH_4SCN added. Indicate any changes as 1 mL or 2 mL increments of NH_4SCN are added.

	Mixture			
NH_4SCN added	Ratio SCN^-/Fe^{3+}	Color hue	Color Intensity (+; ++; or +++)	Solubility
0 mL	0			
1 mL	1:1			
2 mL	2:1			
4 mL	4:1			

D-5. Is there any noticeable difference in color or solubility properties when you add one equivalent of thiocyanate ion to the iron(III) ion compared to when you add four equivalents of thiocyanate ion to the iron(III) ion?

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