Formula of an Ionic Compound
Balancing Charges on Ions

Introduction

Atoms of different elements combine with one another to form compounds. The empirical formula of an ionic compound indicates the kinds of atoms that are present in the compound as well as the relative number (ratio) of each kind of atom. Let's investigate how the formula of an ionic compound can be determined experimentally.

Concepts

- Ionic compounds
- Polyatomic ions
- Empirical formula
- Precipitation reaction

Background

An *ionic compound* is composed of ions—atoms or groups of atoms that have a positive or negative charge. Oppositely charged ions arrange themselves into an extended, three-dimensional structure called a crystal lattice. The net attractive forces among oppositely charged ions in the crystal structure are called ionic bonds. Although composed of charged ions, ionic compounds are electrically neutral. The ratio of oppositely charged ions in the crystal structure is such that the positive charge contributed by the cations is equal to or balanced by the negative charge contributed by the anions. There is no net or overall charge in an ionic compound.

The *empirical formula* of an ionic compound indicates the smallest whole number ratio of each type of ion in the crystal structure and is called a formula unit. For example, magnesium chloride has the empirical formula $MgCl_2$. Magnesium cations (Mg^{2+}) and chloride anions (Cl^-) combine in a 1:2 ratio to form the $MgCl_2$ formula unit. The overall charge on ionic compounds is always zero.

Some ions consist of a charged group of covalently bonded atoms. Such ions are called *polyatomic ions*. An example is the nitrate ion (NO_3^-), which contains one nitrogen atom and three oxygen atoms and has an overall charge of –1. In calcium nitrate, calcium (Ca^{2+}) ions combine with nitrate ions in a 1:2 ratio in order to balance the positive and negative charges. The empirical formula for calcium nitrate is $Ca(NO_3)_2$. Parentheses are used around the nitrate ion to show that the subscript "2" pertains to the nitrate ion as a whole.

Many ionic compounds can be prepared in the lab using *precipitation reactions*. When solutions of two ionic compounds are combined, the ions may rearrange to form a new ionic compound that is insoluble in water. An example of this type of reaction is the formation of solid barium sulfate when barium chloride and sodium sulfate are combined in solution (Equation 1a). In Equation 1b, only the ions that form the precipitate are represented. This makes it easier to recognize what happens in the precipitation reaction.

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq) \qquad \textit{Equation 1a}$$

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s) \qquad \textit{Equation 1b}$$

According to the balanced equation for this reaction, barium ions (Ba^{2+}) combine with sulfate ions (SO_4^{2-}) in a 1:1 ratio to form barium sulfate ($BaSO_4$). This ratio can be observed experimentally in the lab by mixing $BaCl_2(aq)$ and $Na_2SO_4(aq)$ solutions containing equal amounts (concentrations) of barium and sulfate ions, respectively. The maximum amount of precipitate will be obtained when equal volumes (a 1:1 ratio) of the two solutions are combined. A similar approach can also be used to determine the formula of an unknown ionic compound.

Experiment Overview

The purpose of this experiment is to determine the empirical formula of an unknown ionic compound. Two solutions containing equal amounts (concentrations) of two reactant ions will be combined in a series of reactions. In each reaction, the total volume of the two solutions will be held constant while the volume ratio of the reactants is varied. The amount of precipitate obtained in each reaction will be measured and plotted against the volume ratio to find the empirical formula of the product.

ab Questions

1. Many common drugstore chemicals are ionic compounds. Write the correct empirical formula for each of the following compounds.

Common name:	Milk of magnesia	Washing soda	Epsom salt
Chemical name:	Magnesium hydroxide	Sodium carbonate	Magnesium sulfate

2. Solutions of iron(III) chloride and sodium hydroxide were mixed in a series of precipitation reactions, as described in this experiment.

   (a) Name the two possible products in this precipitation reaction and predict their empirical formulas.

   (b) Which product is likely to be insoluble in water and precipitate out as a red solid?

   (c) What volume ratio of reactants gave the most precipitate (see Table 1)? Explain.

Table 1.

Test tube	1	2	3	4	5	6	7
$FeCl_3$, 0.1 M, mL	5	10	12	15	17	20	24
NaOH, 0.1 M, mL	55	50	48	45	43	40	36
Volume of precipitate, mL	1	10	14	20	4	1	0

erials

Copper(II) chloride solution, $CuCl_2$, 0.1 M, 6 mL

Sodium phosphate solution, Na_3PO_4, 0.1 M, 6 mL

Marking pen or wax pencil

Metric ruler, marked in millimeters

Pipets, Beral-type, 2

Stirring rod or wood splints

Test tubes, small, 7

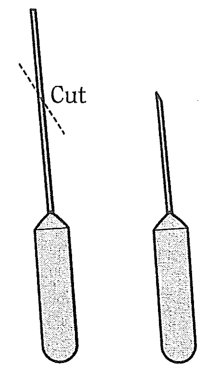
Test tube rack or 24-well reaction plate

See the
Teachi
a discu
observ
of FeCl
(Pre-L
The an
shown
conver
this re
stratio
to lab.
be car
gradu

Safety Precautions

Copper(II) chloride and sodium phosphate solutions are skin and eye irritants and are slightly toxic by ingestion. 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



1. Label seven small test tubes #1–7 with a marking pen and place them in a test tube rack or in a 24-well reaction plate.
2. Cut the stems of two Beral-type pipets at a 45° angle about 5 cm from the bulb, as shown here.
3. Fill one pipet with 0.1 M copper(II) chloride solution and record the color of the solution in the data table.
4. Carefully add the appropriate number of drops of copper(II) chloride solution to each test tube #1–7, as shown in Table 2. *Note:* Exact volumes are very important—hold the pipet vertically to obtain uniform size drops.
5. Fill the second pipet with 0.1 M sodium phosphate solution and record the color of the solution in the data table.
6. Carefully add the appropriate number of drops of sodium phosphate solution to each test tube, as shown in Table 2.

Table 2.

Test tube	**1**	**2**	**3**	**4**	**5**	**6**	**7**
$CuCl_2$, 0.1 M (drops)	3	6	12	15	18	24	27
Na_3PO_4, 0.1 M (drops)	27	24	18	15	12	6	3

7. Use a *clean* stirring rod or wood splint to stir each reaction mixture in test tubes #1–7. Let the tubes sit undisturbed for 10–15 minutes to allow the precipitates to settle.
8. During this time, determine the volume (drop) ratio of copper(II) chloride and sodium phosphate solutions in each test tube. Write this ratio in the data table. *Example:* In test tube #1, 3 drops of $CuCl_2$ and 27 drops of Na_3PO_4 correspond to a 1:9 ratio of $CuCl_2$:Na_3PO_4.
9. After the precipitates have settled, observe the appearance of the products *(both the solid and the solution)*. Record the observations in the data table in the space provided. Be as detailed as possible.
10. Use a metric ruler to measure the height of the precipitate in millimeters in each test tube. Read from the top of the solid material to the bottom center of the test tube. Record each height in mm in the data table.
11. Dispose of the contents of the test tubes as directed by your instructor.

Name: ____________________

Class/Lab Period: ____________________

Formula of an Ionic Compound

Data Table

Color of $CuCl_2$ Solution		Color of Na_3PO_4 Solution	
Appearance of Products			

Precipitation Reactions							
Test Tube	1	2	3	4	5	6	7
Volume Ratio* (Drops $CuCl_2$:Drops Na_3PO_4)	1:9						
Height of Precipitate (mm)							

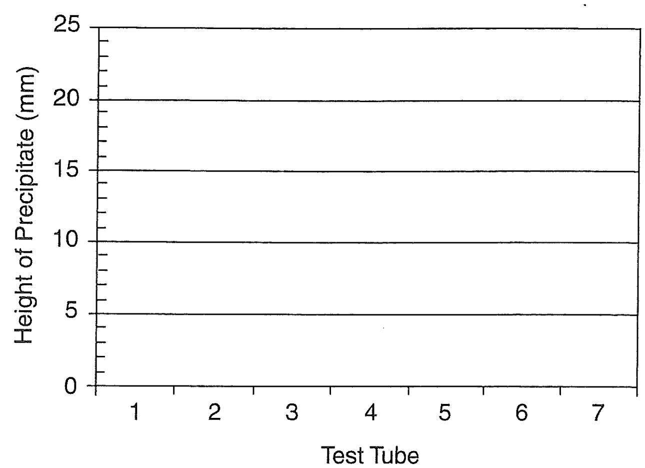
*Reduce the volume ratio to the simplest whole-number ratio.

Post-Lab Questions

1. (a) Name the two possible products in the precipitation reaction of copper(II) chloride with sodium phosphate. Use the charges on the ions to predict the empirical formulas of the products.

   (b) Based on common knowledge, which product is likely to be insoluble in water and to precipitate from solution?

2. Complete the following bar graph to show the height of precipitate in each test tube.



3. Which test tube had the greatest amount of precipitate? Does this result agree with the prediction made in Question #1 concerning the empirical formula of the product? Explain.

4. Write a balanced chemical equation for the precipitation reaction of copper(II) chloride and sodium phosphate. Include abbreviations for the physical state of each reactant and product, using (aq) for aqueous solution, (s) for solid, (l) for liquid, and (g) for gas.

5. (a) Which test tubes showed evidence of unreacted Cu^{2+} ions in the supernatant when the reaction was complete? Explain why unreacted Cu^{2+} ions were present in these tubes based on the volume ratio of solutions used.

5. (b) How could you tell that all of the Cu^{2+} ions had reacted in a particular test tube? Which test tubes showed such evidence? Explain, based on the volume ratio of solutions used.

6. What was the total number of drops of solution in each test tube? Why was it necessary to keep the total volume of reactants constant in each test tube?

7. *(Optional)* Does the *height* of precipitate in each test tube accurately reflect the *amount* of precipitate in each case? *Hint:* Compare the shape of a test tube to that of a graduated cylinder. What effect does this error have on the conclusions reached in this experiment?