Experiment 7 - Thermal Decomposition of Sodium Bicarbonate

Introduction

Decomposition reactions involve a single substance splitting into several substances. Many ionic compounds containing polyatomic ions can decompose when heated. In this lab we will decompose ordinary baking soda, which is sodium bicarbonate. When sodium bicarbonate ($NaHCO_3$) is heated above 110 °C (but not heated to "red heat") both H_2O and CO_2 are produced, and a white solid residue is left when the decomposition is complete. Among possible reactions, the two following <u>unbalanced</u> reactions seem the ones that are most likely to explain these facts:

$$\underset{\text{sodium bicarbonate}}{NaHCO_3(s)} \rightarrow \underset{\text{sodium carbonate}}{Na_2CO_3(s)} + \underset{\text{water vapor}}{H_2O(g)} + \underset{\text{carbon dioxide}}{CO_2(g)} \qquad \textbf{Rxn 1}$$

$$NaHCO_3(s) \rightarrow \underset{\text{sodium oxide}}{Na_2O(s)} + H_2O(g) + CO_2(g) \qquad \textbf{Rxn 2}$$

You are to choose whether **Rxn 1** or **Rxn 2** is most supported by your results. Keep in mind that in many chemical processes, multiple reactions occur in competition, so that the reaction that happens fastest will produce the most product before the reactants run out. First, you will heat a weighed sample of sodium bicarbonate in a crucible and determine the mass of solid product. The mass of the solid product ("residue") will suggest whether the product is more likely Na_2CO_3 or Na_2O.

For additional confirmation, you will try a few chemical tests to identify the presence of carbonate ions. To understand these chemical tests, it will be useful to keep in mind the **anhydride** pattern of reactivity. Anhydride means "without water"; anhydrides are oxide compounds that can undergo combination reactions with water. These reactions are often reversible under the right circumstances. There are **acid anhydrides** (non-metal oxides that produce an acid when combined with water) and **basic anhydrides** (metal oxides that produce a base when combined with water). Both appear in this lab. Carbon dioxide can undergo a reversible anhydride combination with water, making carbonic acid, as follows

$$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq) \qquad \textbf{Rxn 3}$$

The reverse of this reaction produces fizzing when acid is added to carbonate salts. The acid restores the carbonic acid (a weak acid, so it is mostly not dissociated), which can then release the carbon dioxide. Thus, if you observe fizzing on adding acid to your product, carbonate or bicarbonate ions are likely present.

In contrast, if the product is sodium oxide, it can still react with acid. In this case, when acid is added, two reactions occur very quickly (probably in competition with each other). One is a basic anhydride reaction, in which sodium oxide reacts with water to make sodium hydroxide:

$$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq) \qquad \textbf{Rxn 4}$$

The second is a standard acid-base reaction between hydroxide ions and the acid. Both of these reactions release heat, so the mixture will probably get hot.

Safety Precautions:
- Wear your safety goggles.
- Use care when handling HCl(aq). If any HCl splashes on your skin or clothes, rinse it off immediately with plenty of water.

Waste Disposal:
- The solid waste may be discarded in one of the regular garbage cans.
- Liquid waste containing barium should be placed in the waste bottle in the hood. Other liquid waste can be washed down the sink. If you have difficulty cleaning the tube with precipitate, adding a drop or two of acid will help it dissolve.

Procedure

Decomposition Reaction
1. Obtain a clean, dry crucible that has no cracks. (Ideally, use a copper crucible.)
2. Place the crucible in a clay triangle, on a ring attached to a ring stand. Make certain that the crucible hangs securely in the triangle; ask for help if it seems shaky. Put the cover on the crucible.
3. Heat the crucible and cover for 10 minutes with a high flame, until the bottom of the crucible glows red hot. Allow the crucible to cool in the clay triangle. (This step is to clean out the crucible.)
4. When they have reached room temperature, weigh the crucible and cover. It's a good idea to weigh them separately in case the cover breaks.
5. Add sodium bicarbonate to the crucible until it is 3/4 full. Weigh the crucible, sample, and cover. Obtain the mass of the sample by subtraction.
6. Using a medium-sized flame, heat the sample for 15 minutes. Do NOT heat the sample until it is red hot (turn down the flame if it seems too hot). Cool it in place as before, and then weigh it. (While you are waiting for the sample to heat and cool, you may start on the calculations and chemical tests.)
7. Heat again, this time for 10 minutes, then cool and weigh as before. If the crucible, cover, and sample have not changed mass during this heating step, we can be reasonably confident that the sodium bicarbonate has been completely decomposed (i.e. no more H_2O or CO_2 is being produced.)
8. If the crucible plus sample weighs less after this second heating than it did after the first heating, repeat the heating/cooling/weighing sequence until the mass is relatively constant (usually three times is sufficient to achieve a mass loss of less than 10 mg).

Chemical Tests
Carbonate salts such as $Na_2CO_3(s)$ react with acids to form CO_2 gas. If the solid fizzes when some HCl solution is added to it, this suggests that the solid is a carbonate or bicarbonate. However, Na_2O will also react vigorously with acid, because it is a base, so it may not be easy to distinguish in this way.

9. While you are waiting for your reaction to heat or cool, make the following combinations. For each, record observations, and write a chemical equation that is consistent with what you observe. Make sure to include phases in your reactions. Before adding HCl in Reactions A and B, add approximately 1–2 mL of water to a pea-sized amount of the solid reactant and stir briefly.
   A. $MgO(s) + HCl(aq)$
   B. $Na_2CO_3(s) + HCl(aq)$
   C. $NaOH(aq) + Ba(OH)_2(aq)$
   D. $Na_2CO_3(aq) + Ba(OH)_2(aq)$
10. After all of your heatings and weighings are done, place a bit of your product and a small amount of water in a test tube and add a few drops of 6 M hydrochloric acid. Do you observe fizzing or change in temperature? Record your observations. Does it seem more similar to Rxn A or B in Step 9 above?
11. As a second test for carbonate ion, dissolve a bit of your product completely in water in another test tube. Then add a few drops of barium hydroxide solution. What do you observe? Is it more similar to Rxn C or D in Step 9 above? If you believe a reaction happened, write the molecular and net ionic equation. What does this prove, if anything?

Calculations

1. Balance Rxns 1 and 2 from the Introduction.
2. Starting with the mass of sodium bicarbonate that you used, calculate the mass of Na_2CO_3 that should be produced in theory, assuming that reaction 1 is the one that actually occurs. Show your work.
3. Starting with the mass of sodium bicarbonate that you started with, calculate the mass of Na_2O that should be produced in theory, assuming that reaction 2 is the one that actually occurs. Show your work.
4. Compare your observed mass of product with those predicted in step 2 of the calculations for $Na_2O_{(s)}$ and for $Na_2CO_{3(s)}$. On the basis of these answers, which product was actually formed?
5. Calculate the percent yield of the reaction. Does the result make sense? Explain.
6. Write molecular and net ionic equations for each of the reactions you observed in steps 9–11.

Name: Section:

Experiment 7 Pre-Lab Sheet

1. (1 pt) Can you tell the difference between the reactant and the product by looking? Why or why not?

2. (1 pt) How can you tell when the decomposition reaction is complete?

3. (3 pts) Complete these sequential reactions, including phases:

$Na_2O(s) + H_2O(l) \rightarrow$ _______ $+ Ba(OH)_2(aq) \rightarrow$

$Na_2CO_3(s) \xrightarrow[H_2O]{}$ _______ $+ Ba(OH)_2(aq) \rightarrow$

$Na_2CO_3(s) \xrightarrow[H_2O]{}$ _______ $+ 2HCl(aq) \rightarrow$

4. (2 pt) In the $Ba(OH)_2$ tests shown in the previous problem, what difference would you expect to observe? Why would it be a problem if you don't dissolve the decomposition product completely before adding barium hydroxide?

5. (3 pts) Sample calculation: If you perform a similar experiment decomposing $Cu(OH)_2$ according to the following equation, starting with 6.831 g of $Cu(OH)_2$, how much CuO should you get, in grams?

$$Cu(OH)_2(s) \rightarrow CuO(s) + H_2O(g)$$