THE CANADA SCIENCE FAIR JOURNAL

ARTICLE

Chemistry

Experimental Investigation to Evaluate the Stoichiometric Ratio of the Reaction of Hydrochloric Acid and Sodium Carbonate

Written By: Zeyus Spenta
Co-Authors: Jeffery Lee and Dmitri Folomeyev
Edited by: Sabina Yasmin
Layout By: Jonathan Leung

The concept of stoichiometry, which is about relative quantities of the reactants and products of a balanced chemical reaction, is a very important term in chemistry.[1] The concept of mole ratios is a key tenet of stoichiometry.

Introduction

A mole ratio is the ratio of the number of moles relative to one another in a balanced chemical equation. It can be illustrated by the coefficient of each product or reactant in a balanced chemical equation. Thus “[a] balanced chemical equation shows us the numerical relationships between each of the species involved in the chemical change” [1]. Mole ratio is related to mass, and it is also an important topic in stoichiometry. As the moles in a balanced chemical equation can be converted to mass (in grams), it's possible to convert the moles ‘out’ (product) and moles ‘in’ (reactant) into mass. In a balanced chemical equation, the mass of its reactants will be equal to the mass of the products. This is because atom can neither created nor destroyed in a chemical reaction, and its only chemical bonds are broken and formed through a chemical reaction. Thus, the total number of atoms of each element taking part in the reaction will be equal to the total number of atoms of the same element in the product.

The purpose of this study is to establish the validity of stoichiometric ratios as a predictive mechanism in chemistry and connect the knowledge of reaction stoichiometry to experimentation by ascertaining whether or not stoichiometric ratios can be determined in a laboratory setting. In a balanced chemical equation, a certain mass (in grams) of a reactant can be converted into the number of moles. Thereafter, by utilizing stoichiometric ratios, it is possible to determine the number of moles that each of the other reactants and products in the balanced chemical equation will be comprised of.

To evaluate the stoichiometric ratios in a laboratory setting, the following acid-based reaction was chosen.

$$2\mathrm{HCl(aq)} + \mathrm{Na_2CO_3(s)} \rightarrow 2\mathrm{NaCl(aq)} + \mathrm{H_2O(l)} + \mathrm{CO_2(g)}$$

There were several reasons for choosing this reaction. First, the acid (HCl) is a clear aqueous solution, and the base (Na2CO3) is a white powder. It is predicted that the product of the experiment will be salt, water, and CO2. The product salt should be dissolved in water, resulting in a clear and colourless solution. So, it would be very easy to detect the end point of reaction. Again, the product CO2 would naturally evaporate, and it's easy to remove water by heating. Thus, it’s possible to obtain pure NaCl. If the experiment is properly performed, the actual mass of NaCl produced would be close to the calculated value and it would be possible to experimentally prove the validity of stoichiometric ratio.

Materials and Methods:

Materials:

- 1.00 M HCl (in a plastic dropper bottle) (Innovating Science)
- Solid Na2CO3 (Wintersun Chemical)
- 1 x 250 mL beaker
- Plastic weighing boat
- Non-permanent marker
- Digital scale (to 0.001 g precision)
- Heating plate
- Laboratory oven

Methods:

1. First, a clean and dry 250 mL beaker was taken, and its mass was measured and recorded.
2. By using a plastic weighing boat, 1.128 g Na2CO3 (base) was measured, and then transferred into the beaker. The actual mass of Na2CO3 was recorded.
3. Then acid (1.00 M HCl) was added to the beaker containing Na2CO3 dropwise from the dropper bottle until all the Na2CO3 completely reacted with the HCl.
4. Then the beaker was placed onto a hot plate and gently warmed for 3-4 minutes to vent off any excess HCl.
5. The beaker was transferred to an oven and left there overnight to ensure that all the water evaporated from the solution leaving solid NaCl.
6. Next day, the mass of the beaker containing NaCl was measured and recorded. The amount of produced NaCl was determined by subtracting the previously measured mass of the empty beaker from the beaker with the NaCl.

Results and discussion:

Observations:

- The HCl was clear and colourless and was in aqueous form.
- The Na2CO3 was white in colour and was in a solid, powder form.
- During the reaction of HCl and Na2CO3 in the beaker CO2 was produced. CO2 is gas, so escaped from the beaker and the contents of the beaker began to bubble and fizz.
- When an added drop of HCl did not result in the contents of the beaker producing more bubbles and solid Na2CO3 was no longer visible, it was assumed that all the Na2CO3 had fully reacted with the HCl.
- After the completion of the reaction, the contents of the beaker were clear and colourless, and it was composed of NaCl in aqueous form, H2O in liquid form, excess HCl in aqueous form.
- While the contents of the beaker were being gently warmed on a hot plate, condensation appeared on the beaker and the excess HCl being vented off in gaseous form was white in colour.
- After heating overnight, the beaker was composed of finely-ground, flaky NaCl powder that was white in colour and was quite firmly attached to the bottom of the beaker.

Data Table:

Table 1 highlights values calculated throughout the experiment.

Table 1: Recorded values for mass before and after NaCl reaction

Mass of empty 250 mL beaker	118.659 g
Mass of Na_2CO_3	1.128 g
Mass of 250 mL beaker with NaCl after reaction and drying	119.643 g
Mass of NaCl produced	0.984 g

Calculations:

Firstly, the number of moles in the amount of Na2CO3 taken was determined. Then by utilizing the stoichiometric ratios in the balanced equation, the expected number of moles of the NaCl that could be produced was determined. Then using molar mass, the amount of NaCl that could be produced was determined. To determine whether stoichiometric ratios may be experimentally ascertained, the comparison of the mole ratio between the moles of the reacted Na2CO3 and the expected moles of the produced NaCl, with the moles of the reacted Na2CO3 and the actual moles of the produced NaCl used. The number of actual moles produced versus the number of expected moles produced should be very similar.

Determination of the moles of Na_2CO_3 reacted:

The molar mass of $NaCO_3$ = {2(22.990) +1(12.011) +3(15.999)} g/mol =105.988 g/mol

$$\text{Number of moles of } Na_2CO_3 \text{ reacted} = 1.128\ \cancel{g} \times \left(\frac{1\ mol}{105.988\ \cancel{g}}\right) = 0.01064 \text{ mol}$$

Estimation of the moles of NaCl produced based upon how much Na_2CO_3 reacted:

$$\text{Moles of NaCl produced} = 0.01064\ \cancel{mol\ Na_2CO_3} \times \left(\frac{2\ mol\ NaCl}{\cancel{1\ mol\ Na_2CO_3}}\right) = 0.02128 \text{ mol}$$

Calculation of the expected mass of NaCl produced:

The molar mass of NaCl = {1(22.990) + 1(35.453)} g/mol = 58.443 g/mol

$$\text{Calculated mass of crystallized NaCl} = 0.02128\ \cancel{mol}\ NaCl \times \left(\frac{58.443\ g}{1\ \cancel{mol}}\right) = 1.232 \text{ g}$$

Calculation of the percent error:

Calculated mass of crystallized NaCl = 1.232 g

Actual mass of crystallized NaCl = 0.984 g

$$\text{Percent Error} = \left(\frac{Expected\ Number - Actual\ Number}{Actual\ Number}\right) \times 100 = \left(\frac{0.984 - 1.232}{0.984}\right) \times 100 = -25.2\ \%$$

Figure 1: Stoichiometry Calculations

According to calculations done in Figure 1 we can see that the percent error of the actual mass of NaCl produced was -25.2% of the expected mass of NaCl produced.

Sources of Error:

Systematic:

A systematic error that may have taken place during the experiment was that when the reactants of HCl and Na2CO3 were added to the beaker and moved around while the experiment's procedure was being conducted, microscopic splatter took place (this means that some particles within the beaker would have 'escaped' its confines). There were multiple points in the experiment where microscopic splatter could have occurred, such as when the beakers were being moved onto the heating plate and into the laboratory oven. A loss of matter in the beaker would therefore lead the actual mass of the sample to erroneously be lower than it should have been had the error not occurred. Thus, the percent error between the actual mass of the sample vs. the expected mass of the sample (which does not account for microscopic splatter) would be greater as the two values would be further apart. For instance, without microscopic splatter, the actual mass of the NaCl sample may have been 1.103 g in comparison to the 1.244 g theoretically calculated. As such, the difference in the actual vs. expected values in this case would be smaller than the difference between the 0.984 g (actual mass) and 1.244 g (expected mass) which was obtained through the experiment's procedure.

Another systematic error might be caused due to the hygroscopic nature of Na2CO3. Since the Na2CO3 sample was laying out in an open container before the experiment began and was further exposed to the air while it was being measured, some water molecules in the air may have attached to the particles of the Na_2CO_3. Eventually, the mass of the Na_2CO_3 sample increased. So, the actual mass of Na_2CO_3 to react with HCl might be less than the measured value resulting in the formation of less amount of NaCl than expected. However, through the procedure, those excess water molecules would be evaporated as the sample was placed in a laboratory oven. Since the expected mass of the NaCl was calculated based on the measured mass Na_2CO_3 sample, the inclusion of the mass of the water molecules with the Na_2CO_3 sample result in the actual mass being smaller than the expected mass to a certain degree based on how many water molecules attached to the Na_2CO_3 sample (the expected mass' calculations were thus bloated, while the actual mass' calculations were not bloated by this systematic error occurring).

Random Error

One source of random error that might occur during the experiment was the precision of the scales being utilized. The scales used can only go to .001g of precision, therefore meaning that the measurement could have been off by a degree of ±0.0005g.

This could mean that if the weighing scale happened to be more precise, perhaps a greater mass of Na_2CO_3 than illustrated by the weighing scale could have been obtained, thus increasing the expected mass of NaCl due to stoichiometric ratios. Or, if the more precise value of Na_2CO_3 happened to be less than what the scale measured, then the expected mass of NaCl would have been smaller. However, due to the precision of the equipment, it was unable to figure out which scenario might have occurred. Furthermore, when the actual mass of the NaCl sample produced was measured, the precision of the value obtained was limited by the scale. A more precise value of the mass of the NaCl would have made the percent error much smaller.

While the random error displayed in the experiment may have occurred, it would play a very tiny role in shaping the experiment. The precision of the equipment alone certainly does not account for the actual mass of NaCl produced being 25.2% less than the expected mass. As such, the contribution of systematic error to the results obtained through the experiment must be considered.

Conclusion

In conclusion, stoichiometric ratios can be experimentally derived save for systematic and random sources of error. This was illustrated in an experimental setting where Na_2CO_3 was reacted with HCl and produced NaCl. It was expected to produce 1.244 g of NaCl using stoichiometric mole ratios. However, this was not the case as the actual grams of NaCl was 25.2% less than the expected number of grams.

To identify why the mass of the NaCl sample obtained through the experiment did not align with the expected mass of NaCl, a few possible systematic and random errors were analysed. In terms of the systematic errors microscopic splatter and the hygroscopic nature of the Na_2CO_3 sample could contribute to the fact of the determination of lower mass of NaCl. The precision of the weighing scales used is considered one source of random error. However, this error could have only minorly altered the result of the experiment. Since the results illustrated the actual mass of the produced NaCl sample was less than the mass of NaCl expected to be produced, it can be assumed that the influence of the systematic errors as a whole resulted in lowering the actual mass of the NaCl sample in comparison to the NaCl sample's expected mass.

However, it can be concluded that the stoichiometric ratios in a balanced chemical equation can be theoretically determined, but in a laboratory setting these ratios may be inaccurate due to experimental error. If all these errors in the experiment could eliminated, it is possible to obtain the theoretical stoichiometric ratios in a balanced chemical equation.

References

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