Determination of the Amount of Acid Neutralized by an Antacid Tablet by Back Titration

Goal and Overview

A traditional antacid tablet contains one or more base that can react with the hydrochloric acid, HCl(aq), in gastric juice. Because the tablet is often only sparingly soluble in neutral solutions, it will be dissolved in a known amount of excess HCl. A portion of the total HCl is neutralized by the tablet; the remainder will be quantified by back titration with NaOH(aq).

Objectives and Science Skills

Understand and explain standardization as it applies to acidic and basic solutions used as reagents in an experiment.
Define back titration and explain why it is used.
Determine the average acid neutralizing capacity of an antacid and its associated standard deviation based on statistical treatment of data from multiple titration trials.
Quantitatively and qualitatively compare experimental results with theoretical values and evaluate factors that may contribute to observed deviations.

Suggested Review and External Reading

Reference materials and textbook sections covering acid-base neutralization reactions, acid-base titration, and pH indicators

Introduction

The primary acidic component in gastric juice is hydrochloric acid, HCl. In a healthy stomach, digestion functions properly when the pH is about 3. Occasionally excess acid is produced, leading to internal discomfort and/or damage. Traditional antacid tablets (versus proton-pump inhibitors) contain bases such as calcium carbonate, $CaCO_3$, and magnesium hydroxide, $Mg(OH)_2$, that can neutralize stomach acid.

$$CaCO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$
$$Mg(OH)_2(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + 2\ H_2O(l)$$

The base content in an antacid tablet is generally given in milligrams on the product label. Common regular strength tablets may contain 550 mg $CaCO_3$ and 110 mg $Mg(OH)_2$; extra strength 675 mg $CaCO_3$ and 135 mg $Mg(OH)_2$. The actual acid-neutralizing capacity of a tablet is checked by experiment and can be compared to the theoretical value.

Acid-base titration is frequently used for the quantitative analysis of acids and bases in solution. During the titration process, precise volumes of a standardized solution of base or acid at known molarity (the "titrant") are added to a solution containing an unknown amount of acid or base (the "analyte"). The equivalence point in the titration occurs when the stoichiometric amount of base or acid has been added and the acid or base in the analyte is exactly neutralized.

Monitoring the progress of the neutralization reaction can be accomplished by adding a pH indicator to the analyte. The indicator should show a color change over a fairly narrow pH range that includes the approximate pH expected at the equivalence point in the titration. The color change provides a visual signal that you have reached the equivalence point (technically, this is the endpoint, which is often slightly past the equivalence point but is within experimental tolerance limits).

Many antacid tablets are not very soluble in water. If you put a tablet in water and try to titrate it with HCl directly, too much HCl is added before the tablet dissolves. To overcome this problem, back titration can be used. The tablet is dissolved in a known quantity of excess HCl. The antacid tablet neutralizes part of the total HCl; the remainder is titrated with NaOH.

Equipment List

Pan balance, weigh paper
Safety goggles and appropriate attire
Buret, clamp, and stand
Volumetric pipet and pipetter
Deionized water
Erlenmeyer flasks (125 mL or 250 mL) and stoppers (#5 or #6)
Small beakers
Labeling tape and markers

Procedure

Please pay attention to your TA's demonstration and ask questions.

You need data for four good titration trials. You may be instructed to write the tablet masses and equivalence volumes (V_{eq}) on the board or to otherwise share your data.

Parts 1 and 2. Standardization of NaOH and HCl – A separate sheet of instructions will be provided if you are required to standardize the NaOH and HCl that you will use in Part 3.

Part 1. Standardized NaOH

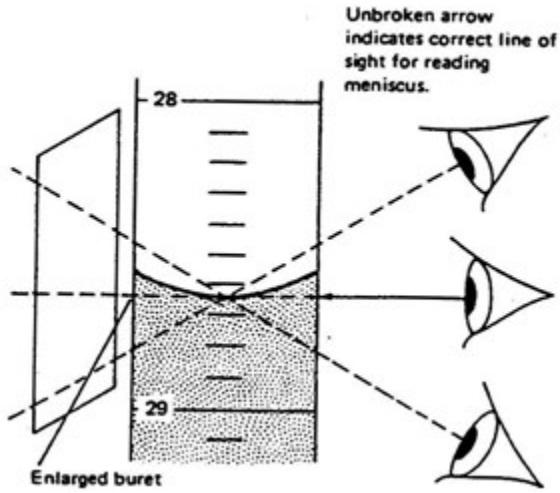
1. Record the molarity of the NaOH to four decimal places.

2. Carefully take about 50 mL in a small labeled beaker. This is enough to fill your buret one time. Once you see how much NaOH a single titration requires, you can take more if you need it. Please minimize the amount you waste (you cannot pour excess back into the reagent bottle).

3. Carefully fill your buret (see Caution below). Record the initial volume reading, V_{in}, at the bottom of the meniscus to 2 decimal places. V_{in} does NOT need to be 0.00 mL.

Caution

Make sure that the buret is secure and straight in the buret clamp that is fastened to the stand. There should always be a labeled container under the buret large enough to capture the liquid contents.
Do not pour chemicals above eye level. You can carefully set the stand on the floor long enough to fill the buret. Use a funnel; have a labeled waste beaker beneath the buret; and, make sure that the stopcock is closed.
Carefully set the buret stand back on your lab bench. Allow a small amount of NaOH to pass through the tip into the waste beaker. Record the initial volume, V_{in}, to two decimal places.



V = 28.65 mL; readings are usually ±0.05 mL

(on a line, ##.#0 mL, or halfway between two lines, ##.#5 mL).

Part 2. Standardized HCl

4. Record the molarity of the HCl to four decimal places.

5. Carefully take about 100 mL in a small labeled beaker. You will perform four titration trials, using 25.00 mL HCl per titration.

Part 3. Determination of the Amount of HCl Neutralized by an Antacid Tablet

6. Record the amounts of active ingredients in the antacid tablet listed on the container (number of mg $CaCO_3$ and/or number of mg $Mg(OH)_2$). Take four tablets.

7. Record the masses of the tablets on the pan balance to two decimal places. The tablets may have slightly different masses, so keep track of which tablet you are titrating in each trial.

8. Add a tablet to one of four 125-mL Erlenmeyer flasks labeled with that tablet's mass (250-mL flasks are also acceptable).

9. Add about 25 mL deionized water to each flask. This volume does not need to be very precise so using a graduated cylinder is fine.

10. The volume of HCl added to the flask must be as precise as possible. Use your 25.00 mL volumetric pipet and the correct pipetter to add 25.00 mL HCl into each flask.

Caution

Incorrect use of the pipet and pipetter can result in **serious injury**.
Please **do NOT force or jam the pipet** into the pipetter.
If you push too hard, the pipet can break and **cut your hand**.
Use **just enough force** to create a seal. The pipet does not "click" into place.
Hold both the pipet and pipetter at all times.
If you pull any solution into the pipetter, please give it to your TA so that it can be cleaned.

If you have trouble, **ask your TA for help**.

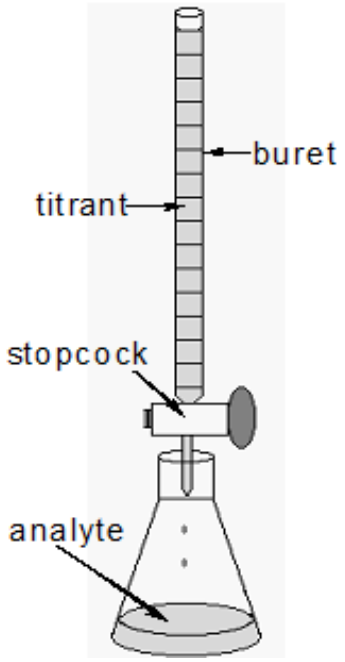
11. Heat the solutions gently to a near boil for about 5 minutes to drive off any CO_2 generated by the reaction of a carbonate base with HCl. The tablet should dissolve completely (or to a great extent). DO NOT PUT MORE THAN TWO FLASKS on the hot plate at one time.

Caution

The flasks contain acidic solutions. Please handle them carefully.
Make sure the flasks are stable on the hot plate before turning on the heat (no more than two flasks at one time).
Once the flasks have been exposed to heat, assume everything is hot. Do not touch the hot flasks.
Keep an eye on the solutions while you perform other tasks. Do not let the solution boil vigorously or spatter out of the flask.

12. Allow the solution to cool so that it is not warm to the touch and stopper it until you are ready to titrate.

13. Add 2-3 drops of bromothymol blue (BTB) indicator to the cool analyte solution. In solutions with pH < 6.0, BTB appears yellow; pH > 7.6, blue. Place a light-colored sheet of paper or paper towel under your flask.



$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

14. Make sure that you have recorded V_{in} from the buret to two decimal places.

15. Add NaOH from the buret to the flask in **~1 mL** portions while swirling the analyte solution. When you see a temporary blue that disappears readily with swirling, decrease the volume of NaOH per addition to **~0.5 mL**. When the blue persists slightly longer, go to

dropwise additions. At the endpoint, the color should be **pale lime green**. One drop of NaOH can cause this. If the color goes from **yellow to blue, stop**. You have passed green. Do not add any more NaOH.

16. Record V_f to two decimal places.

17. Once you know approximately how much NaOH is required to reach the endpoint, you can titrate your second, third, and fourth solutions more quickly by adding a larger volume of NaOH before switching to the ~0.5 mL and dropwise additions close to the endpoint.

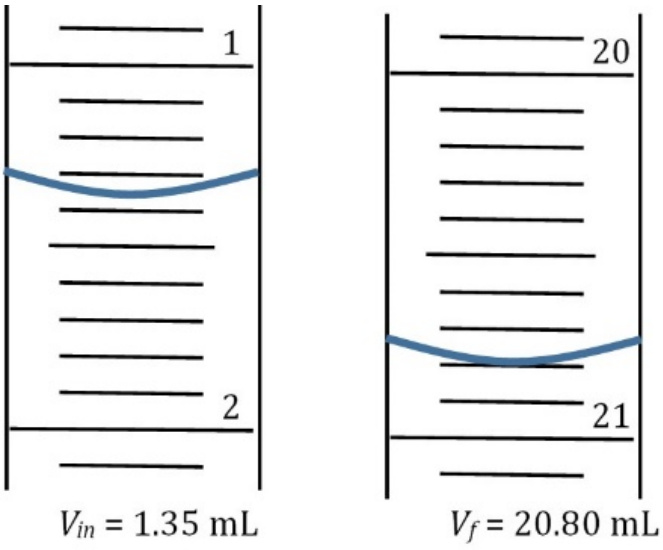
18. Dispose of liquid waste in the waste containers in the back hood. Wash all glassware well (tap water and soap, thorough tap water flush, and final deionized water rinse); please be very careful. Clean your work area, making sure that it looks as good or better than when you began. Return any equipment that you borrowed clean.

Please use your experimental data, not the values shown in examples.

19. Calculate the total number of moles of HCl used in each trial (*use your molarity; do not assume that it was exactly 1.0000 M).

$$n_{\text{HCl total}} = V_{HCl} \times M_{HCl} = (0.02500\ \text{L}) \times \left(\frac{1.0000\ \text{mol}}{1\ \text{L}}\right)^{*} = 0.02500\ \text{mol}$$

15. Calculate the volume of NaOH that was required to reach the equivalence point, V_{eq}, in each trial to two decimal places. Assume $V_{endpoint} = V_{eq} = V_f - V_{in}$.



V_{in} = 1.35 mL V_f = 20.80 mL

$V_{NaOH\ added}$ = 20.80 – 1.35 mL =19.45 mL

20. Calculate the moles of HCl neutralized by NaOH in each trial (*use your molarity; do not assume it was exactly 0.5000 *M*)

$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

$$n_{\text{HCl by titration}} = n_{OH^- \text{ at } V_{eq}} = V_{eq} \times M_{NaOH} = (V_{eq} \text{ in L}) \times \left(\frac{0.5000 \text{ mol}}{1 \text{ L}}\right)^* = x_{trial\ \#} \text{ mol}$$

21. Calculate the average and standard deviation for the tablet masses, V_{HCl}, $V_{NaOH\ at\ eq}$, $n_{HCl\ total}$, and $n_{HCl\ titration}$. Recall that σ should be rounded to one significant figure and the associated average to that decimal place.

22. Calculate the number of moles of HCl neutralized by each tablet. Find the unrounded average and standard deviation and then round to the correct number of significant figures.

$$n_{\text{HCl by tablet}} = n_{\text{HCl total}} - n_{\text{HCl titrtation}}$$

23. Calculate the theoretical number of moles of HCl each tablet should neutralize to three significant figures based on the milligrams of active ingredients indicated on the product label (assume the mg values are good to the ones place).

$CaCO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$
$Mg(OH)_2(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + 2\ H_2O(l)$

24. Quantitatively compare (to three significant figures) the experimental acid neutralizing power of an average tablet relative to the theoretical value. Also, find the average moles of HCl neutralized per gram tablet.

$$\text{actual vs. expected HCl neutralized} = \frac{\text{experimental average } n_{\text{HCl by tablet}}}{\text{theoretical } n_{\text{HCl by tablet}}} \times 100\%$$

$$\text{average HCl neutralized per gram tablet} = \frac{\text{average } n_{\text{HCl by tablet}}}{\text{average tablet mass}}$$

Please follow your TA's instructions for cleanup and waste disposal. Return any equipment and glassware borrowed clean.

Results / Sample Calculations

Complete the online inlab or write a lab report as directed by your TA.

Moles of HCl – total, neutralized by titration, and neutralized by tablet (including averages and standard deviations)
Actual vs. expected moles of HCl neutralized by the tablet
Average moles of HCl neutralized per gram tablet

Discussion Questions and Review Topics

What did you find and how did you do it?
How similar were the tablet in terms of the amount of acid neutralized?
How did your experimental findings compare to the expected value?
What were the major experimental sources of error ("human error" is not an acceptable answer), and what could you do to minimize these?

Supplement – Standardization Instructions

Standardization of NaOH (if required)

Determine the concentration of the base, NaOH, by titrating a known mass of the monoprotic acid, KHP, to neutral (the equivalence point). The molar mass of KHP is 204.23 g/mol, and it has one acidic hydrogen per molecule.

1. Mass out approximately 1.000 g potassium acid phthalate (KHP). About 10.00 mL NaOH should be used in the titrations. The NaOH solution's concentration is about 0.5 M. The molar mass of KHP is 204.23 g/mol, and it has one acidic hydrogen per molecule.

2. Put the KHP into 50–100 mL deionized water in a 250-mL titrating flask. It does not need to dissolve completely, and you don't need to know how much water is in the flask. The KHP is functioning as a strong acid, and it will dissolve as it is titrated. You can warm the water to aid the dissolution if needed.

3. Add a few drops of BTB indicator to the titration flask.

4. Record the initial volume of NaOH from the buret and then titrate the KHP solution to the endpoint. Record the final volume from the buret and calculate V_{eq}.

Standardization of HCl (if required)

To determine the precise molarity of the HCl solution, titrate it with the NaOH to the endpoint; use BTB as the indicator unless instructed otherwise.

1. Add about 25 mL deionized water to a 125-mL flask and then volumetrically pipet 10.00 mL of HCl into the flask Erlenmeyer flask.

2. Record the initial volume of NaOH from the buret and then titrate the HCl to the endpoint. Calculate V_{eq}.

3. Repeat until you can get reproducible results (at least three consistent trials).