GOVERNMENT DEGREE COLLEGE, ETURNAGARAM,

MULUGU (Dist.)

JIGNASA 2019-20

A Student Study Project

Conducted Under the Jurisdiction of Commissioner of college Education, Telangana State

DEPARTMENT OF CHEMISTRY

Topic

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

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Introduction:

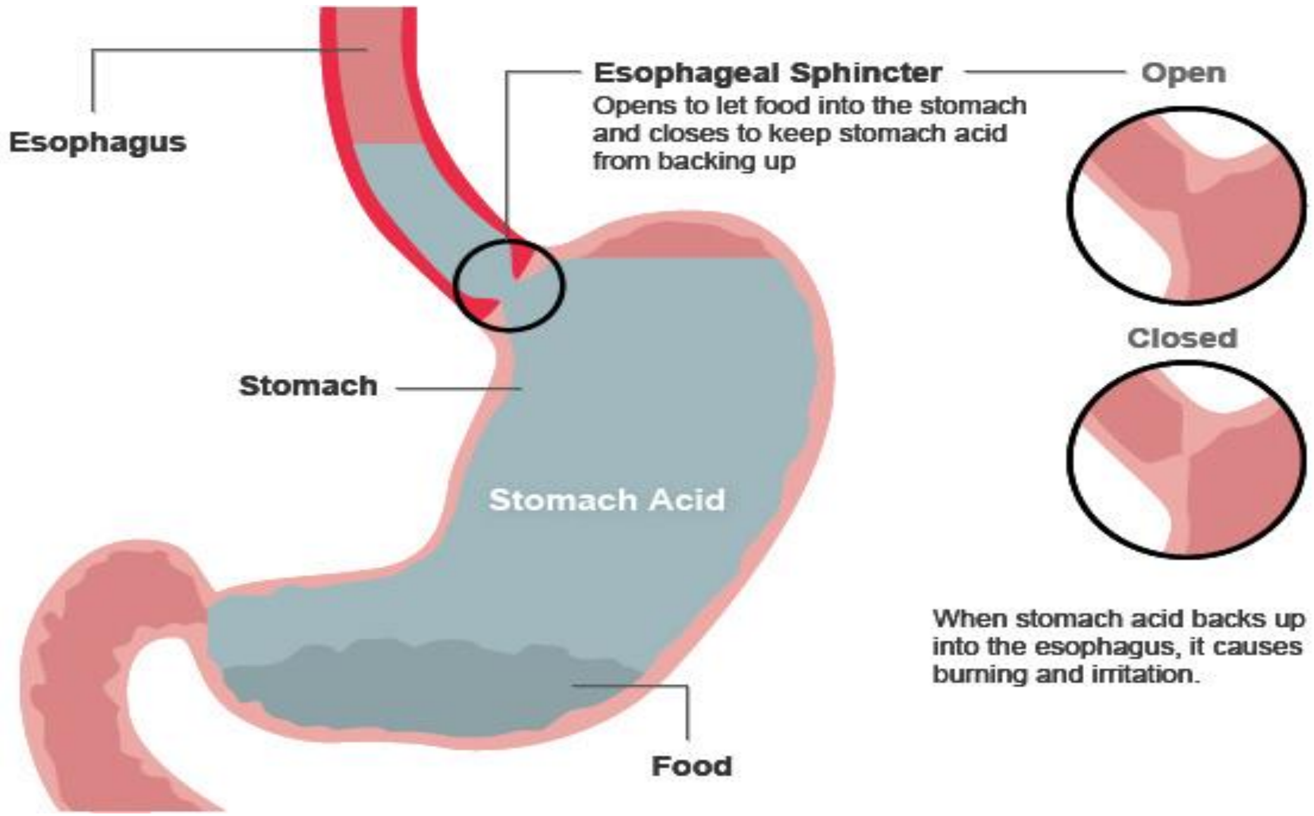
Antacids are bases that react stoichiometrically with acid. The number of moles of acid that can be neutralized by a single tablet of a commercial antacid will be determined by back titration. To do the experiment, an antacid tablet will be dissolved in a known excess amount of acid. The resulting solution will be acidic because the tablet did not provide enough moles of base to completely neutralize the acid. The solution will be titrated with base of known concentration to determine the amount of acid not neutralized by the tablet. To find the number of moles of acid neutralized by the tablet, the number of moles of acid neutralized in the titration is subtracted from the moles of acid in the initial solution.

Objectives of the study:

- 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 neutralized 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.

Side effects

- Versions with magnesium may cause diarrhea, and brands with calcium or aluminium may cause constipation and rarely, long-term use may cause kidney stones. Long-term use of versions with aluminium may increase the risk for getting osteoporosis.


Esophagus
Esophageal Sphincter
Opens to let food into the stomach and closes to keep stomach acid from backing up
Open
Closed
Stomach
Stomach Acid
When stomach acid backs up into the esophagus, it causes burning and irritation.
Food


Tablets
I.P. 150 mg
Zinetac 150 mg
Ranitidine
Tablets
I.P. 150 mg
Zinetac 150 mg


Orange Flavour
ANTACID ANTIGAS TABLETS
DIGENE
B.NO.070113D7 MFD.07/2019 EXP.06/2022
M.R.P. Rs.16.50 INCL. OF ALL TAXES


Original
Antacid Antigas
Chewable Tablets

Review of related literature:

Acid-base reactions and the acidity (or basicity) of solutions are extremely important in a number of different contexts-industrial, environmental, biological, etc. The quantitative analysis of acidic or basic solutions can be performed by titration. In a titration, one solution of known concentration is used to determine the concentration of another solution by monitoring their reaction.

To determine the amount of base in an actual tablet, ideally you would dissolve it in water and titrate with acid. In most titrations, solutions of the acid and the base are used. This is not an option here because $CaCO_3$ is quite insoluble in water. By the time tablet completely dissolves, you will have added too much acid.

To overcome this problem, the antacid tablet is dissolved in a known amount of excess acid: the excess acid is neutralized with more base.

One factor to consider: since the tablet contains a carbonate, the neutralization reaction produces carbon dioxide. Because CO_2 dissolves in water to produce carbonic acid, H_2CO_3, it can cause your results to be off. You will drive off the CO_2 by heating the solution just below boiling for about 5 minutes to alleviate this problem.

Another factor to consider: acidic and basic solutions are generally colorless. How can you tell when you have reached the endpoint of the titration? At the endpoint, the amounts of strong acid (e.g., H^+) and strong base (e.g., OH^-) are equal. The pH changes dramatically with addition of more acid or base.

An acid-base indicator gives a visual indication of the acidity or base city of a solution. The indicator is usually an organic dye that behaves as a weak acid or a weak base. The indicator's color depends on whether it is in the dissociated or un dissociated from (which depends on the pH of the solution): Hln = H^+ (aq) + In^-

Hln is the un dissociated from that is dominant at lower pH levels; In^- is the conjugate base (remains after dissociation) that is dominant at higher pH levels. HIn has one color and In^- another. The equilibrium constant for this weak acid is:

$$K_a = \frac{[H^+][In^-]}{[HIn]}$$

Research Methodology

Apparatus: Burette 10 ml. pipette, conical flask, beakers and fennel etc.

Chemicals required:

1. Standard (0.05 M) Na_2CO_3 solution
2. HCl solution
3. NaOH solution
4. Antacid

Indicator:

1. Methyl Orange
2. Phenolphthalein

Principle: Estimation of alkali in antacid by using standard HCl is a back-titration method. In this method excess of HCl solution (Known volume) is added to antacid. Some amount of HCl reads with alkali in antacid and the unreached HCl remains present in the antacid solution. The volume of unreached HCl remains present in the antacid solution. The volume of unreached HCl is determined by using standard NaOH solution. From the volume of NaOH consumed the volume of HCl reached with alkali content in the antacid and then the amount of alkali present in antacid are determined.

Chemical equations:

1. Standardization of HCl:

$$Na_2CO_3 + 2HCl \longrightarrow 2NaCl + H_2O + CO_2 \uparrow$$

1 Mole Na_2CO_3 = 2 Moles of HCl

2. Standardization of NaOH:

$$HCl + NaOH \longrightarrow NaCl + H_2O$$

3. Estimation of alkali in Antacid:

i) $\underset{\text{(antacid)}}{OH^-} + \underset{\text{(excess)}}{HCl} \longrightarrow Cl^- + H_2O + \underset{\text{(Unreached)}}{HCl}$

ii) $HCl + NaOH \longrightarrow NaCl + H_2O$ (Un reached)

Procedure:

1. Preparation of standard (0.05 M) Na_2Co_3 solution :

About 0.55 gr of Na_2Co_3 is taken in a weighing bottle and weighed. It is W_1, gr. After Na_2Co_3 is transferred in to 100 ml. volumetric flask, the empty weighing bottle is weighed. It is w_2 gr.
(w_1-w_2) gr gives the weight of Na_2Co_3 transferred in to the volumetric flask. The substance Na_2Co_3 is dissolved in a little amount of distilled water and then the solution is made up to the mark. The solution is thoroughly mixed. From the weight of Na_2Co_3 dissolved, the morality of the solution is calculated.

. Molarity of Na2Co3 solution: (w_1-w_2) /GMW x V (ml) {GMW of Na_2Co_3 = 106 gr}

Calculations:

Weight of the weighing bottle + Na_2CO_3 (w_1) gr = 8.105 gr

wt of the weighing bottle (after Na_2CO_3 is transferred) (w_2) gr = 7.575 gr

.: wt of Na_2co_3 (w_1-w_2) gr = 8.105-7.575 = 0.053 gr

.: Molarity of the Na_2CO3 solution = (w_1-w_2) / GMW x V (l) =0.053 x 10/106 = 0.05M

{GMW of Na_2CO_3 = 106}

2. Standardization of HCl:

The burette is rinsed with distilled water and then with a little HCl solution. Now the burette is filled with HCl and the initial reading is noted.

Well cleaned 10 ml pipette is rinsed with standard Na_2Co_3 solution. Then 10 ml Na_2Co_3 is pipette out into a conical flask and 2 drops of methyl orange indicator are added. The solution gets pale yellow color. Now the Na_2Co_3 solution is titrated against the HCl solution until the color of solution turns to pale pink. This is the end point. The titration is stopped and the final reading is noted.
The titrations are repeated until incurrent values are obtained. From the values obtained, the molarity of HCl is calculated.

S. No.	Volume of the Standard Na_2CO_3 solution (Pipette reading)	Burette readings (in ml)		Volume HCl consumed (in ml) F.R- I. R
		Initial	Final	
1	10 ml	0	10	10
2	10 ml	10	20.2	10.2
3	10ml	0	10	10

$$Na_2CO_3 + 2\,HCl \longrightarrow 2\,NaCl + H_2O + CO_2 \uparrow$$

1 mole Na_2CO_3 =2 moles of HCl

$$V_1M_1/n_1 = V_2M_2/n_2$$

Na_2Co_3 values

Molarity (M_1) = 0.05M

Volume (PR) (V_1) = 10ml

No. of moles (n_1) = 1

HCl Values

Molarity (M_2) =?

Volume (BR) (V_2) = 10 ml

No. of moles (n_2) = 2

.: $M_2 = V_1M_1n_2/V_2n1$ = 0.05 x 10 x 2 /10 = 0.12 M

.: Molarity of HCl (M_2) = 0.12 M

3. Standardization of NaOH solution (Blank titration):

The burette is filled with standard HCl solution. 10 ml pipette is rinsed with the given NaOH and then 10 ml. NaOH is pipette out in to a conical flask. 2 drops of phenolphthalein indicator are added to it. Then the solution gets pink color. Now NaOH solution is titrated against HCl. At the end point the pink color of the solution is disappeared. The final reading of the burette is noted. It is ‘ x’ ml.

The titrations are repeated until concurrent values are obtained. From the volume ‘ x’ ml. (i.e., volume of HCl neutralized by 10 ml of NaOH) we can calculate the volume of NaOH requires for 10 ml of HCl (which will be added to the antacid) . It is V_1 ml.

S. No	**Volume of the Standard NaOH solution (in ml) (PR)**	**Burette readings (in ml)**		**Volume HCl consumed (in ml) F.R-IR (x ml)**
		Initial	Final	
1	10 ml	0	9	9
2	10 ml	9	18.5	9.5
3	10ml	0	9	9

‘9 ‘ml. of HCl is neutralized by 10 ml. of NaOH.

10 ml of HCl is neutralized by V_1 ml. of NaOH

.: V_1= 10 x 10/ x = 100/9 = 11.5 ml

.: Volume of NaOH required for 10 ml. of HCl (V_1) ml = 11.5 ml

4. Estimation of Alkali content in antacid:

The burette is filled with standard NaOH solution and the initial reading is noted.

5 ml of antacid suspension is taken in to a 100 ml. standard flask and make it up to the mark with distilled water.

10 ml. of this antacid solution is pipette out into a conical flask and 10 ml of standard HCl is added. The conical flask is covered and heated to 70 o c on a hot water bath for about 10 minutes and cool. [Some amount of HCl reacts with alkali OH^- in the antacid and the un reacted HCl remains present in the solution]

Then the cooled solution (containing un reacted or excess HCl) is back titrated with standard NaOH solution using phenolphthalein indicator. At the end point the color of the solution is turned to light pink color. The final reading is noted.

From this value, the volume of NaOH required for HCl which reacted with alkali in antacid. From the volume of NaOH, the volume of HCl reacted with alkali in antacid and then the amount of alkali present in the given 100 ml of antacid can be determined.

S. No	**Volume of antacid +HCl solution**	**Burette readings (in ml)**		**Volume NaOH consumed (for un reacted HCl in antacid) F.R-I. R (V_2) ml**
		Initial	Final	
1	10 ml	0	10	10
2	10 ml	10	21	11
3	10ml	0	10	10

a) Volume of NaOH required for un reacted HCl in the Antacid solution (V_2) ml = 10

Volume of NaOH required for 10 ml of HCl (V_1) ml = 11.5 ml

.: volume of NaOH requires for HCl which reacted with alkaline Antacid (V_1-V_2) ml = 11.5-10 = 1.5

b) 10 ml of NaOH (Black titration) (x ml) 9 ml of HCl
(V_1-V_2) ml of NaOH = (V_1-V_2) x/10 = 1.5 x 9/10 = 2.02 (y) ml of HCl

c) $M_1V_1/n_1 = M_2V_2/n_2$

Antacid Values	**HCl Values**
Molarity (M_1) =?	Molarity (M_2) = 0.12
Volume (PR) = V_1 = 10ml	Volume (V_2) = 2.02 (y) ml
No. of moles = n_1 = 1	No. of moles (n_2) = 1

.: $M_1 = M_2V_2n1/n_2 \times V_1$

= 0.12 x 2.02 / 10

.: Molarity of antacid solution (M_1) = 0.041 M

.: The amount of alkali OH –in antacid = M_1x 17/10 = gr/ 100ml

= 0.041 x 17/10 = 0.069 gr/ 100ml

Results:

- Antacids are a class of drugs used to treat conditions caused by the acid that is produced by the stomach. The stomach naturally secretes an acid called hydrochloric acid that helps to break down proteins.
- This acid causes the contents of the stomach to be acidic in nature, with a pH level of 2 or 3 when acid secretion is active.
- The stomach, duodenum, and esophagus are protected from acid by several protective mechanisms.
- Antacids reduce acidity by neutralizing (counteracting) acid, reducing the acidity in the stomach, and reducing the amount of acid that is refluxed into the esophagus or emptied into the duodenum.
- Antacids also work by inhibiting the activity of pepsin, a digestive enzyme produced in the stomach that is active only in an acid environment and, like acid, is believed to be injurious to the lining of the stomach, duodenum, and esophagus.

Conclusion:

- In conclusion, the synthetic antacids had the greatest effect on the pH difference of the HCl, but at the same time they were too potent in the fact that they brought the pH level too high to be in a healthy range for a stomach.

- A healthy pH for an antacid to bring the stomach acid to be is about 2 or 3, so if the pH gets higher than 3.0, it only secretes more acid to keep the pH below 3.0.

- When heartburn or acid reflux medication interferes with stomach acid by raising the pH above 3.0, the stomach is no longer functioning properly.

- It then creates chemical combinations that are not usable by the body which then create more problems for the rest of the digestive track because it then has to work harder to break the food particles down.

- when that happens, fermentation occurs which leads to gas and bloating and more discomfort. As well, Galveston for example failed to bring the pH to a level high enough to be considered a natural environment for the stomach.