Alex Iacob

CHMG 146 Section 12

Date of Analysis 4/2/2021, 4/9/2021, 4/16/2021

Lab 5: Alkalinity

Abstract

Alkalinity is often described as the capacity of water to resist acidification in terms of concentration. This concentration corresponds to the amount of acid which is added as a titrant. In this case, the milligrams per liter calcium concentrate is used. To determine this, three separate experiments are used. The first focuses on standardizing a sulfuric acid solution via titration with carbonate to find the average molarity of the sulfuric acid. The second focuses on determining the alkalinities of an unknown substance using titration and visual cues. The third focuses on determining the alkalinities of the same solution but using pH probes to measure the rate of change in pH then graphing to find the endpoints. The unknown sample was determined to be a solution of CO_3^{2-} and HCO_3^-. The total alkalinity of this solution was calculated to be 2833.33 mg of $CaCO_3$ / L using Part 2's method and 3125 mg of $CaCO_3$ / L using Part 3's method.

Experimental

Part 1 of the lab was dedicated to standardizing a solution of sulfuric acid. This was done via the process of titration. First, 1 gram of Na_2CO_3 was measured and added to a 250mL flask and 50mL of distilled water to dissolve the salt. Another 50mL of distilled water was added to create the "Standard Na_2CO_3 Solution, " then its molarity was calculated using *Equation 1*.

Equation 1: Used to find the molarity of Standard Na_2CO_3 Solution

$$\frac{1.03 \; grams \; of \; Na_2CO_3}{0.200 \; L} * \frac{99.96 \; grams \; of \; Na_2CO_3}{100 \; grams} * \frac{1 \; mol \; of \; Na_2CO_3}{105.96 \; grams \; of \; Na_2CO_3} = 0.0486 \; M \; of \; Na_2CO_3$$

Next, the titration buret was conditioned by passing through sulfuric acid then was filled to the 0 mL line. Afterward, 20mL of the Standard Na_2CO_3 Solution was added to a flask to begin the process of titration. To that flask, 20mL of distilled water and 3-4 drops of phenolphthalein were added. Before titration, this solution was a vibrant pink color. Next, the value for the first endpoint, EP_1, was calculated using *Equation 2*.

Equation 2: Estimating the amount of sulfuric acid to reach EP$_1$ (theoretical value)

$$0.02\ L\ *\ \frac{0.0486\ M}{1\ L\ of\ soln}\ *\ \frac{1\ mol\ of\ H^+}{1\ mol\ of\ CO_3^{2-}}\ *\ \frac{1\ mol\ of\ H_2SO_4}{2\ mols\ of\ H^+}\ *\ \frac{1\ L\ of\ soln}{0.05\ moles\ of\ H_2SO_4}\ =\ 9.7\ L\ of\ H_2SO_4$$

Once the theoretical value was attained, the solution was titrated until it became a transparent color. The amount of sulfuric acid required to reach that clear color was recorded, then 3-4 drops of a mixed indicator were added to create a blue-green color. The flask was placed on a heat plate then cooled down in ice water. The solution was then titrated until the discernable blue-green color faded away. The amount of sulfuric acid required to achieve this transparency was noted as EP_2. The molarity of the sulfuric acid was calculated using *Equation 3*. This process has been repeated a total of three times to get three different trials. The results for this first part are recorded in *Results Table 1*.

Equation 3: Used to find the molarity of sulfuric acid

$$\frac{20\ mL\ of\ standard\ soln}{EP1\ or\ EP2}\ *\ \frac{0.0486\ mol\ of\ CO_3^{2-}}{1\ L\ standard\ soln}\ *\ \frac{1\ mol\ of\ H^+}{1\ mol\ of\ CO_3^{2-}}\ *\ \frac{1\ mol\ of\ H_2SO_4}{2\ mols\ of\ H^+} =\ M\ of\ H_2SO_4\ from\ EP1\ or\ EP$$

Part 2 of the experiment was to determine the alkalinity of an unknown. This experiment begins with the selection of one of many solutions: solution ε for this experiment. 20 mL of the unknown solution, 20mL of distilled water, and 3-4 drops of phenolphthalein were added to a total of three flasks. Once added, the solution became a vibrant pink, which indicated a need for titration. The amount of mL required to titrate the solution was recorded as EP_1. Once titrated to

a clear solution, 3-4 drops of a mixed indicator were added. Once added, the solution became a

blue-green color. This indicated that another titration process needed to occur. The flask was

heated up, then cooled down and titrated until the blue-green color began to fade. The mL of

sulfuric acid needed to titrate the solution was recorded as *EP₂*. Once every trial was completed,

the alkalinities were calculated using *Equation 4*, with the *A* value being represented from

Results Table 3.

Equation 4: Used to calculate the alkalinities of the solution with given endpoint values

$$\frac{A\ mL\ of\ H_2SO_4}{20\ mL\ sample\ soln} * \frac{0.05\ mol\ of\ CaCO_3}{1\ L\ of\ H_2SO_4} * \frac{1\ mol\ of\ CaCO_3}{1\ mol\ of\ H_2SO_4} * \frac{100\ g\ of\ CaCO_3}{1\ mol\ of\ CaCO_3} * \frac{1\ mg\ of\ CaCO_3}{0.001\ g\ of\ CaCO_3} = mg\ of\ CaCO_3/L$$

Part 3 of this experiment was to estimate the potentiometric titration of an unknown. To

begin, a pH meter was calibrated, then 20 mL of solution ε and 50mL of distilled water were

added to a beaker. The pH of the solution was measured in 0.5mL increments of the addition of

H_2SO_4. The second set of points was calculated to determine the endpoints to determine the

alkalinities using *Equation 3* and *Results Table 2*.

Results and Discussion

Table 1: The average molarity of sulfuric acid and trials for Experiment 1

	Trial 1		**Trial 2**		**Trial 3**	
	EP_1	EP_2	EP_1	EP_2	EP_1	EP_2
Total vol of Na_2CO_3	20 mL	20 mL	20 mL	20 mL	20 mL	20 mL
Init vol of H_2SO_4	0 mL	10.6 mL	10 mL	20.8 mL	32 mL	35 mL
Final vol of H_2SO_4	10.6 mL	20.6 mL	20.8 mL	30.4 mL	42.2 mL	44.6 mL
Total vol of H_2SO_4	10.6 mL	10 mL	10.8 mL	9.6 mL	10.2 mL	9.6 mL
Molarity of H_2SO_4	0.0485 M	0.0486 M	0.0450 M	0.0506 M	0.0477 M	0.0506 M
Average M of H_2SO_4	0.0485 M					
Standard Deviation	0.0019					
% RSD	3.912 %					

Table 2: Calculations for A values to be used in Alkalinity

	Total Alk	Pheno Alk	Carb Alk	Bicarb Alk	Hydrox Alk
A Value	$EP_1 + EP_2$	EP_1	$2 * EP_1$	$EP_2 - EP_1$	0

Table 3: Titration data for solution ε

	Trial 1		Trial 2		Trial 3	
	EP_1	EP_2	EP_1	EP_2	EP_1	EP_2
Vol of the unknown sample	20 mL	20 mL	20 mL	20 mL	20 mL	20 mL
Init vol of H_2SO_4	0 mL	4.9 mL	9 mL	13.6 mL	12 mL	16.5 mL
Final vol of H_2SO_4	4.9 mL	12.1 mL	13.6 mL	19.9 mL	16.5 mL	23 mL
Total vol of H_2SO_4	4.9 mL	7.2 mL	4.6 mL	4.5 mL	4.5 mL	6.5 mL
Total Alkalinity	3025		2725		2725	
Phenolphthalein Alkalinity	1225		1150		1150	
Carbonate Alkalinity	2450		2300		2300	
Bicarbonate Alkalinity	575		425		425	
Hydroxide Alkalinity	0		0		0	

Table 4: Averages of Alkalinities from Table 2 for convenience

Total Alkalinity	2833.33 mg of $CaCO_3$ / L
Phenolphthalein Alkalinity	1166.67 mg of $CaCO_3$ / L
Carbonate Alkalinity	2333.33 mg of $CaCO_3$ / L
Bicarbonate Alkalinity	500 mg of $CaCO_3$ / L
Hydroxide Alkalinity	0 mg of $CaCO_3$ / L

Table 5: Plot points for pH of the solution

Line 1		Line 2	
H_2SO_4 mL	pH	H_2SO_4 mL avg	$-\Delta pH / \Delta mL$
0	9.81	0.25	0.36
0.5	9.63	0.75	0.38
1	9.44	1.25	0.34
1.5	9.27	1.75	0.5
2	9.02	2.25	0.84
2.5	8.6	2.75	1.78
3	7.71	3.25	1.36
3.5	7.03	3.75	0.52
4	6.77	4.25	0.44
4.5	6.55	4.75	0.4
5	6.35	5.25	0.34
5.5	6.18	5.75	0.38
6	5.99	6.25	0.22
6.5	5.88	6.75	0.28
7	5.74	7.25	0.38
7.5	5.55	7.75	0.3
8	5.4	8.25	0.42
8.5	5.19	8.75	0.76
9	4.81	9.25	1.86
9.5	3.88	9.75	2.22
10	2.77	10.25	1.92

Graph 1: The potentiometric titration of solution ε made from graphing Table 5
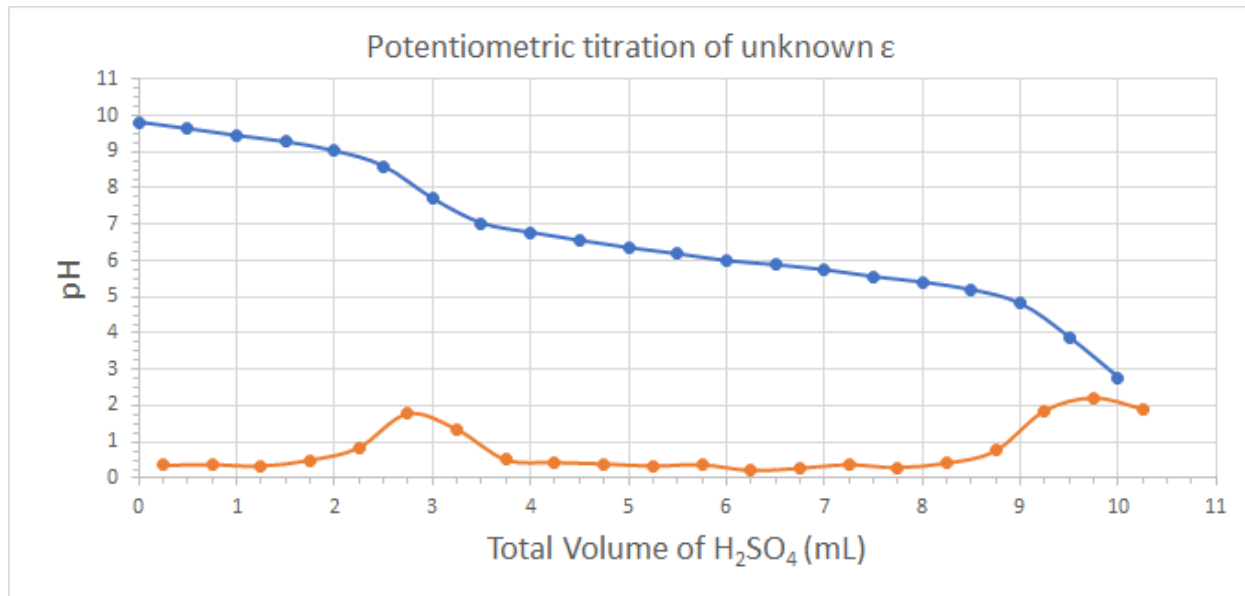
Table 6: Average alkalinities and endpoint values from Graph 1

Endpoint 1	2.75 mL
Endpoint 2	9.75 mL
Total Alkalinity	3125 mg of $CaCO_3$ / L
Phenolphthalein Alkalinity	687.5 mg of $CaCO_3$ / L
Carbonate Alkalinity	1375 mg of $CaCO_3$ / L
Bicarbonate Alkalinity	1750 mg of $CaCO_3$ / L
Hydroxide Alkalinity	0 mg of $CaCO_3$ / L

The main differences between the data from *Table 4* and *Table 6* reside in the method used to attain the values. The data from Table 6 would be more accurate in terms of accuracy because the solution's pH was being taken in shorter intervals rather than being calculated from experimental data. Possible sources of error would include titrating too fast, pH probe not being calibrated correctly, and transfer errors.