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Chemistry 112 – Section 201
Dates of Experiment: March 8 and March 22

Experiment 10 – Acid-base Titrations: Part A – Analysis of vinegar and Part B – Analysis of a Carbonate/Bicarbonate mixture

Unknown Number: __796__

Abstract

The purpose of this experiment was to standardize two titrant solutions in order to identify the cheapest brand of vinegar and the molarities of $NaHCO_3$ and Na_2CO_3 within an unknown solution. To come to these conclusions, the principles and methods of acid-base titrations were used. The experiment explored concepts such as the creation of stock solutions, standardizing a secondary standard, reactions between acids and bases, and the titration of these acids and bases, along with observing their change in pH. Initially, the titrant (NaOH, a base) was prepared in a stock solution and standardized with the use of 100% pure KHP and the indicator phenolphthalein. This was done with several titrations to determine the exact molarity of the stock solution prepared. With this standardization, the NaOH was used to volumetrically analyze the concentration of two brands of vinegar. By determining these concentration values and comparing them to the price and total volume of each vinegar, the better priced vinegar was determined.

With the use of a standardized solution of HCl, an unknown solution of sodium bicarbonate and sodium carbonate was able to be analyzed. Through the use of two separate sets of titrations, both the total alkalinity and the individual molar amount of $NaHCO_3$ were determined. Success in this experiment was contingent upon accurately standardizing the NaOH and HCl solutions, determining accurate concentrations and the best buy for each of the vinegar solutions, and determining the accurate molarities of sodium bicarbonate and sodium carbonate.

Procedure

To prepare a one-liter sample of 0.1 M NaOH, a more concentrated sample was taken and diluted. The smaller amount of concentrated NaOH was poured into a 1-liter plastic bottle, and distilled water was poured to the level of the bottle's shoulder and then was mixed thoroughly.

After mixing the stock solution, the analytical balances were used to weigh about 201.5 mg of KHP into each of three 200 mL tall-form beakers, recording the exact weight of KHP in each. This was followed by adding about 25 mL of distilled water, followed by a stir bar and 2 drops of phenolphthalein solution. The flask was then put onto the stirrer.

A 50 mL buret was rinsed with some of the stock NaOH solution and then filled with the NaOH solution. At this point, the pH sensor was set up. The Lab-quest was connected to the computer via a USB port, and the pH was connected to channel 1 of the Lab-quest. The pH probe was removed from the storage solution and immediately, the bulb was rinsed with water from the squirt bottle. Quickly, the pH probe was secured to a ring stand using a clamp, and then the pH probe was lowered into the solution that would soon be titrated. The bulb was not allowed to touch anything in the beaker, as it was very fragile. The pH was recorded before adding titrant, at

the endpoint, and after the titration was completed, and the initial volume reading of the buret was recorded before any titration began.

Once the initial measurements were made, the first solution was titrated until a consistent faint pink endpoint was reached. This titration was conducted by holding the stopcock and slowly letting out the titrant drop by drop, going very slowly as the endpoint was neared. The final volume reading on the buret and the pH at the endpoint were recorded. The titration was repeated with the other two KHP samples until consistent results were obtained.

Next, the vinegar samples were analyzed. Two vinegar sample were collected, noting the brand name, the original volume of the bottle, and its price. 2 mL of the first vinegar brand was pipetted into each beaker, and then water and 2 drops of phenolphthalein were added. The first solution was titrated, and the initial volume, final volume, and pH at the endpoint were recorded Two more titrations were repeated with the same brand of vinegar. Finally, three more titrations were conducted with the second brand of vinegar, noting all of the same measurements for the brand and during the titration.

After the analysis of the two vinegar samples was complete, a stock solution of HCl was produced and standardized. This was done by first preparing 250 mL of 0.1 M HCl by diluting 25 mL 1 M HCl to 250 mL in a small plastic bottle with distilled water. The solution was shaken to mix.

Three of the 200 mL tall-form beakers were obtained, and about 104.8 mg of Na_2CO_3 were added to each, recording the exact amount added. A stir bar was then added to each. About 100 mL of boiled water was obtained in a clean 100 mL graduated cylinder, and about 25 mL of this boiled water was added to each of the beakers containing Na_2CO_3. The hot plate was set to around 250° C, and two 50 mL burets were cleaned and prepared for titrations by clamping them to the ring stand as before. A funnel was used to pour the two different titrants into each buret. One of the burets was rinsed and filled with the stock solution of HCl, and 2 drops of phenolphthalein was added to each of the beakers containing Na_2CO_3 that were to be titrated. The initial volume on the HCl buret was recorded, and the titration was conducted by slowly allowing the acid solution to drain drop by drop. Once the initial pink color had just disappeared, the endpoint was a little over halfway reached. At this point, 2 drops of bromcresol green was added to the solution, and the titration was continued.

Because of the presence of CO_2 being produced, it was difficult to reach the accurate endpoint. Therefore, the solution was titrated until there was a slight change from original blue color to a greenish tinge. The solution was then boiled, and once this was done, it should have returned to blue. A few more drops of titrant were added to reach the endpoint. The final volume and final pH were recorded. The Vernier temperature probe was used throughout the titration as was previously done for the first nine titrations. Two more good titrations were repeated in the same fashion. This completed the standardization of the HCl solution.

To determine the total alkalinity of an unknown solution, 2 mL of the unknown were pipetted into each of the 200 mL tall form beakers with a volumetric pipet, followed by adding 25 mL of boiled water to each, along with 2 drops of phenolphthalein and a stir bar to each. The three solutions were titrated with the standardized HCl solution. The bromcresol green was added at its appropriate time, once the phenolphthalein endpoint was reached. To determine the amount of $NaHCO_3$, 2 mL of the unknown were again pipetted into each of the three beakers, 15 mL of water was added to each, along with 2 drops of phenolphthalein. The second buret was rinsed and filled with the stock NaOH solution produced previously. About 15 mL of the NaOH solution and about 5 mL of 10% $CaCl_2$ was added to the beakers containing the unknowns. Each

was mixed thoroughly. The solutions were then titrated with HCl until the pink color disappeared. The mixture appeared milky because of the precipitate that formed. The other two beakers were also titrated and the pH was recorded for all of the runs. Once all the titrations had been completed, the NaOH solution was stored, and the rest of the materials were cleaned with distilled water.

Data:

Standardization of NaOH Stock Solution			
	Trial 1	**Trial 2**	**Trial 3**
Mass of KHP	0.2015 g	0.1996 g	0.2038 g
Moles of KHP	9.867×10^{-4} mol	9.774×10^{-4} mol	9.979×10^{-4} mol
pH – Before adding titrant	4.68	4.53	4.59
pH – At endpoint	9.3	8.66	8.5
Initial Volume of NaOH	8.30 mL	27.35 mL	18.61 mL
Final Volume NaOH	18.61 mL	36.03 mL	27.35 mL
Volume NaOH used	10.31 mL	8.68 mL	8.74 mL
Concentration of NaOH	0.09570 M	0.1126 M	0.1142 M
Average Molarity	0.1075 M		
Average Normality	0.1075 N		
Standard Deviation	8.369×10^{-3} M		
Relative Standard Deviation	7.785 %		

Titration of Vinegar Samples						
	Brand 1: Pompeian Gourmet White Wine Vinegar			Brand 2: Great Value Apple Cider Vinegar		
	Trial 4	**Trial 5**	**Trial 6**	**Trial 7**	**Trial 8**	**Trial 9**
Original Volume	475 mL			946 mL		
Original Cost	$1.93			$1.24		
pH – Before adding titrant	3.39	3.63	3.61	3.44	3.67	3.66
pH – Endpoint	7.92	8.28	8.00	8.01	8.04	7.87
Initial Volume of NaOH	00.61 mL	16.38 mL	32.01 mL	01.70 mL	17.49 mL	33.30 mL
Final Volume NaOH	16.38 mL	32.01 mL	47.60 mL	17.49 mL	33.30 mL	49.20 mL
Volume of NaOH used	15.77 mL	15.63 mL	15.59 mL	15.79 mL	15.81 mL	15.90 mL
Moles of Base	1.695×10^{-3} mol	1.680×10^{-3} mol	1.676×10^{-3} mol	1.697×10^{-3} mol	1.700×10^{-3} mol	1.709×10^{-3} mol
Moles of Acid	1.695×10^{-3} mol	1.680×10^{-3} mol	1.676×10^{-3} mol	1.697×10^{-3} mol	1.700×10^{-3} mol	1.709×10^{-3} mol
Average Acid Moles	1.684×10^{-3} mol			1.702×10^{-3} mol		
Acid Molarity	0.8420 M			0.8510 M		

Acid Mass in aliquot taken	0.1021 g	0.1032 g
Grams of Acid per Liter	51.07 g/L	51.61 g/L
Mass of Vinegar	2.010 g	2.010 g
% Acetic Acid	5.080 %	5.134 %
Cost of Vinegar in Moles per Dollar	0.207 moles per dollar	0.649 moles per dollar
Better Buy?	Brand 2: Great Value Apple Cider Vinegar	

Standardization of HCl Stock Solution			
	Trial 1	**Trial 2**	**Trial 3**
Mass of Na_2CO_3	0.1048 g	0.1048 g	0.1052 g
Moles of Na_2CO_3	9.888×10^{-4} mol	9.888×10^{-4} mol	9.925×10^{-4} mol
Initial HCl Volume	0.80 mL	21.70 mL	0.95 mL
Final HCl Volume	21.70 mL; 21.29 to 21.70 mL	42.1 mL; 21.70 to 22.55 mL	21.29 mL; 22.55 to 23.65 mL
Volume HCl Used	21.31 mL	21.25 mL	21.44 mL
Concentration of HCl	0.09280 M	0.09306 M	0.09258 M
Mean Molarity	0.09281 M		
Mean Normality	0.09281 N		
pH – Endpoint	4.23		

Titrations for Total Alkalinity			
	Trial 4	**Trial 5**	**Trial 6**
Volume of Unknown	2 mL	2 mL	2 mL
Initial buret Volume	1.60 mL	21.10 mL	0.61 mL
Final buret Volume	22.70 mL; 21.25 to 22.20 mL	43.5 mL; 22.20 to 23.30 mL	21.25 mL; 23.30 to 24.51 mL
Volume HCl used	22.05 mL	23.50 mL	21.76 mL
Total mmol HCl used	2.046 mmol	2.181 mmol	2.020 mmol
Average HCl used	2.082 mmol		
pH – Endpoint	3.75		

Titrations for $NaHCO_3$			
	Trial 7	**Trial 8**	**Trial 9**
Volume of Unknown	2 mL	2 mL	2 mL
Total Volume of NaOH (at beginning and after addition of HCl)	15.45 mL	15.50 mL	15.50 mL
Total Volume of HCl	5.60 mL	6.50 mL	6.10 mL
Total mmol NaOH used	1.661 mmol	1.666 mmol	1.666 mmol
mmol HCl need to titrate excess NaOH	0.520 mmol	0.603 mmol	0.566 mmol

mmol NaOH used in converting $NaHCO_3$ to Na_2CO_3	1.141 mmol	0.603 mmol	1.100 mmol
mmol $NaHCO_3$ converted to Na_2CO_3	1.141 mmol	0.603 mmol	1.100 mmol
mmol HCl consumed by $NaHCO_3$	1.141 mmol	0.603 mmol	1.100 mmol
mmol HCl consumed by Na_2CO_3	0.941 mmol	1.019 mmol	0.982 mmol
mmol Na_2CO_3	0.471 mmol	0.510 mmol	0.491 mmol
Molarity of $NaHCO_3$	0.5705 M	0.5315 M	0.5500 M
Average Molarity $NaHCO_3$	0.5507 M		
Molarity of Na_2CO_3	0.236 M	0.255 M	0.246 M
Average Molarity Na_2CO_3	0.246 M		

Sample Calculations:

Total Volume of Titrant Used:
Final Volume – Initial Volume = Volume used to titrate

18.61 mL – 8.30 mL = $\boxed{10.31 \text{ mL}}$

Concentration of NaOH (standardization):

Moles KHP = Moles NaOH used to titrate

$\frac{\text{Mol KHP}}{\text{L NaOH}}$ = Concentration (M)

$\frac{9.867 \times 10^{-4} \text{ mol}}{0.01031 \text{ L}}$ = $\boxed{0.09570 \text{ M}}$

Average Normality:

For NaOH, since 1 mol NaOH lends 1 mol OH⁻, Molarity = Normality

0.1075 M NaOH = $\boxed{0.1075 \text{ N NaOH}}$

Standard Deviation:

$\sigma = \sqrt{\frac{1}{N}(\sum_{i=1}^{N}(\text{mean} - \text{each individual})^2)}$, where N is the total amount of individuals in the set.

$$\sigma = \sqrt{\frac{1}{3}((0.09570 - 0.1075)^2 + (0.112 - 0.1075)^2 + (0.1142 - 0.1075)^2)} = \boxed{8.369 \times 10^{-3} \text{ M}}$$

Relative Standard Deviation:

$$\frac{\sigma}{\text{mean}} \times 100\% = \text{Relative Standard Deviation}$$

$$\frac{8.369 \times 10^{-3} \text{ M}}{0.1075 \text{ M}} \times 100\% = \boxed{7.785\%}$$

Number of moles of titrant (based on initial concentration and volume added to analyte):

$$\text{molarity of titrant } (\frac{\text{mol}}{1000 \text{ mL}}) \times \text{volume added (mL)} = \text{moles of titrant used}$$

$$\frac{0.1075 \text{ mol NaOH} \times 15.77 \text{ mL}}{1000 \text{ mL}} = \boxed{1.695 \times 10^{-3} \text{ mol}}$$

Acid Moles: Since the there is a 1:1 stoichiometric relationship, Acid moles = Base moles

Acid Mass in the aliquot taken:

Acid moles (mol) × Molecular Weight of acid (g/mol) = acid mass (grams)

1.684×10^{-3} mol × 60.65 g/mol = $\boxed{0.1021 \text{ g}}$

Grams of Acid per Liter:

Molarity of acid (mol/L) × Molecular Weight (g/mol) = g/L of acid

0.8420 mol/L × 60.65 g/mol = $\boxed{51.07 \text{ g/L}}$

Percentage of Acetic Acid:

$$\frac{\text{grams of acid (g)}}{\text{grams of vinegar (g)}} \times 100\% = \text{Percentage of acetic acid in vinegar}$$

$$\frac{0.1021 \text{ g}}{2.010 \text{ g}} \times 100\% = \boxed{5.080 \%}$$

Cost of each vinegar in moles per dollar:

$$\frac{\text{Total volume of vinegar in bottle (mL)}}{\text{Price of bottle (\$)}} = \text{volume per dollar}$$

$$\text{Molarity of Acid } \left(\frac{\text{mol}}{1000 \text{ mL}}\right) \times \text{volume per dollar (mL)} = \text{Cost (moles per dollar)}$$

$$\frac{475 \text{ mL}}{\$1.93} = 246 \text{ mL per dollar}$$

$\frac{0.8420 \text{ mol}}{1000 \text{ mL}} \times 246 \text{ mL} = \boxed{0.207 \text{ moles per dollar}}$

Standardization of HCl: Similar calculations as that done for standardization of NaOH

Total mmol HCl added to titrate excess NaOH:
HCl molarity (mmol/mL) $\times$ volume HCl added (mL) = mmol HCl

0.09281 (mmol/mL) $\times$ 5.50 mL = $\boxed{0.520 \text{ mmol HCl}}$

Excess NaOH: Equal to total mmol HCl added to titrate it (above calculation), since it is a 1:1 stoichiometric ratio

mmol of NaOH used in converting $NaHCO_3$ to Na_2CO_3:

Total NaOH added to solution (mL) – Excess NaOH (mL) = mmol used to convert

1.661 mmol – 0.520 mmol = $\boxed{1.141 \text{ mmol NaOH}}$

Since there are again 1:1 ratios from the chemical equations (discussed in Discussion section), the millimolar amount of $NaHCO_3$ converted to Na_2CO_3 and mmol HCl consumed by $NaHCO_3$ are also both equal to $\boxed{1.141 \text{ mmol}}$.

mmol HCl consumed by Na_2CO_3:

mmol HCl added in total alkalinity titration (average value) – mmol HCl consumed by $NaHCO_3$ = mmol

2.082 mmol – 1.141 mmol = $\boxed{0.941 \text{ mmol HCl}}$

mmol Na_2CO_3:

$\frac{\text{mmol HCl comsumed by Na2CO3}}{2} = \text{mmol } Na_2CO_3$

$\frac{0.941 \text{ mmol HCl}}{2} = \boxed{0.471 \text{ mmol Na}_2\text{CO}_3}$

Molarity of $NaHCO_3$ or Na_2CO_3:

$\frac{\text{mmol of bicarbonate or carbonate}}{2.000 \text{ mL}} = \text{Molarity (M)}$

$\frac{1.141 \text{ mmol NaHCO3}}{2.000 \text{ mL}} = \boxed{0.5705 \text{ M NaHCO}_3}$
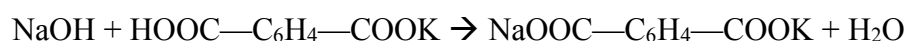
Discussion

One of the many applications of chemistry is the ability to manipulate chemical properties or reactions to analyze certain unknown quantities. One of these analytical processes is called volumetric analysis, in which the concentration of an unknown substance is determined through a process called titration. Titrations utilize one solution of known concentration and volume that is added to an analyte (or the solution being analyzed) of unknown concentration. The solution that is added to the analyte is normally called the titrant. While there are many forms of titrations, this experiment specifically used acids and bases as the titrants and analytes.

While the end goal is to discover a concentration, the procedural goal of a titration is to reach the titration's endpoint, the point at which the desired color change occurs. This is used to indicate when enough volume of titrant has been added to reach the equivalence point of the reaction, the point at which the moles of titrant and analyte have neutralized each other . This is different from the endpoint, but since strong acids and bases are used in the titration process, the pH changes so drastically with incredibly small amounts of titrant added near the equivalence point that the endpoint can be considered to have the same amount of titrant added as the equivalence point. The endpoint is easily attainable because of the use of indicators, which are chemicals that change color at the endpoint, thus fulfilling the function of their name. These are specifically chosen so that their endpoint lies within a small range of the equivalence point of the reaction, thus allowing the moles of titrant needed to neutralize a solution to be accurately determined.
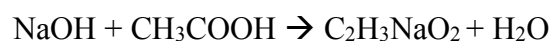
Prior to using a titrant to analyze the concentration of a solution, the titrant must be standardized, meaning that the exact concentration of the titrant itself must be determined. In the case of NaOH, the first titrant used, a 100% pure sample of known concentration of KHP was

used as the primary standard. This purity is a necessity for primary standards; primary standards

are needed because normal solutions are usually slightly impure, so it would be useless to use a

second impure substance to standardize another impure substance. For the indicator of this

titration, phenolphthalein was used. Since NaOH is a strong base and KHP a weak acid, the

equivalence point lies somewhere just above pH = 7 and around pH = 8 or 9. Since

phenolphthalein's range is around pH = 8.3, this is a reasonable choice of indicators, seeing as

the color will change very near the equivalence point.

$$NaOH + HOOC—C_6H_4—COOK \rightarrow NaOOC—C_6H_4—COOK + H_2O$$
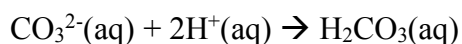
Once the titration was completed and the color had just barely changed, the exact number

of moles of NaOH could be deduced because of the known amount of moles in the KHP (the

weight was measured out exactly). In general, one mole of acid reacts with one mole of base to

create a salt and water, which is considered neutral. Because NaOH and KHP react in a 1:1

stoichiometric relationship, the moles of KHP are equal to the moles of NaOH. All that is left to

find the molarity of the stock solution is to divide by the volume that was added into the analyte.

The value obtained in this experiment was very near 0.1 M (being 0.1075 M), which was the

expected value based on the dilutions conducted in creating the stock.

With this standardized titrant, analysis of unknowns could begin. The first solutions that

were studied were vinegar solutions. The titration process was used to analyze which of two

brands of vinegar would offer the best sale in regards to having the most moles per dollar. To

come to this conclusion, two brands were separately titrated, each using the NaOH. Because

acetic acid is a weak acid, its equivalence point was also around pH = 8 or 9, and thus

phenolphthalein was again a good indicator choice.

$$NaOH + CH_3COOH \rightarrow C_2H_3NaO_2 + H_2O$$

Using the same deductions as with the standardization, the molarity of each of the two acetic acid solutions was determined. Then, by using the price and volume of the original bottles of vinegar, the number of moles of acetic acid per dollar were calculated for each bottle. While the percentages of acetic acid in each vinegar were very similar to each other, the apple cider vinegar brand was undoubtedly the better buy, primarily because the volume (and thus total number of moles by calculation) was much larger for a smaller price than the gourmet vinegar.

For the second set of solutions, which were mixtures of sodium carbonate and sodium bicarbonate, an acidic titrant had to be prepared. This is because the carbonate and bicarbonate ions can act as bases, which means that to neutralize them, an acidic titrant was needed. A stock solution of HCl was prepared by dilution and standardized in a similar fashion to the NaOH stock solution. However, the primary standard used was a known amount of pure Na_2CO_3. There is also another key difference to this second standardization: KHP only needed one OH^- molecule for every KHP molecule to become neutralized, but each Na_2CO_3 molecule must receive two protons to become fully neutralized. Because of this, two endpoints are associated with Na_2CO_3, one that occurs when the Na_2CO_3 has accepted the first proton and one when it has accepted the second and become neutralized. The net equation for this reaction is as follows:
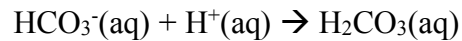
$$CO_3^{2-}(aq) + 2H^+(aq) \rightarrow H_2CO_3(aq)$$

The first endpoint that occurs happens around a pH of 8, so phenolthalein was used again. The second ending occurred in the more acidic range of a pH around 3.5 to 5.5. This is around the range of the bromcresol green indicator.
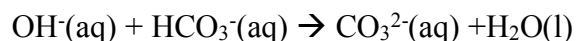
A second difference and complication in titrating the carbonate is that a byproduct of CO_2 is formed during the titration. This dissolved CO_2 actually acts as a buffer for the solution, thus making it difficult to neutralize, since the buffer works to resist the change in the pH. To solve

this problem, the solution was boiled once the endpoint was soon to be reached. With the CO_2

boiled out, the endpoint could be reached normally and without the resistance of a buffer.

With the standardized HCl titrant, an unknown concentration of sodium bicarbonate and

sodium carbonate could be analyzed to determine the concentrations of both the bicarbonate and

carbonate in the solution. To do this, the solution needed to be titrated twice, each time under

different conditions. The first titration was used to determine the total alkalinity of the solution,

which was determined based on the total amount of HCl consumed to neutralize the solution.

$$HCO_3^-(aq) + H^+(aq) \rightarrow H_2CO_3(aq)$$

$$CO_3^{2-}(aq) + 2H^+ \rightarrow H_2CO_3(aq)$$

This equation shows that 1 mole of HCl is consumed for each mole of bicarbonate and 2

moles are consumed for each mole of carbonate. However, unless either the bicarbonate amount

of the carbonate amount is known, nothing more than the total sum of the amounts of

bicarbonate and carbonate can be deduced. This is why the second titration was necessary; it

used the NaOH solution to react with any bicarbonate to produce only carbonate ions, which was

then precipitated out with $CaCl_2$. This method cleverly allows one to analyze only the amount of

bicarbonate within the solution, regardless of how much carbonate is present. By adding NaOH

and $CaCl_2$, the solution was left with excess OH^- ions, and by titrating with HCl, this excess

amount could be determined. Since the total amount of OH^- added was known, the exact amount

that reacted with bicarbonate could be calculated, and by the 1:1 stoichiometric ratio of the

chemical reaction, the amount of bicarbonate in the mixture could as well be determined.

$$OH^-(aq) + HCO_3^-(aq) \rightarrow CO_3^{2-}(aq) + H_2O(l)$$

With this value, the equation relating the HCl consumed in the titration for total alkalinity

could now to be solved for the amount of carbonate, since two of the three pieces were known.

[Total moles HCl consumed] = [moles consumed by HCO_3^-] + [moles consumed by CO_3^{2-}]

Finally, with these molar amounts identified, the molarity of the two compounds could be calculated in the unknown solution in the same fashion as the previous molar calculations.

Because these types of titrations necessitate slight estimations, there is inevitable error present in these procedures. By performing multiple trials and calculating averages, this error can be acceptably minimized, but error could appear at any stage of the analysis (when determining stock concentration or a specific unknown). This incredibly slight error comes about from the difference between the endpoint and equivalence point. Another error can come about if not all of the CO_2 boils out and continues to act as a buffer, causing slightly more of the HCl to be added than what would be indicative of the actual endpoint. Overall, these errors can be minimized to maintain suitable accuracy in the analysis.

Conclusion

Overall, this experiment was a success. The concentrations for the two stock solutions (0.1075 M NaOH, 0.09281 M HCl) were both near the expected values of 0.1 M, and the best purchase for vinegar was determined to be the great value apple cider vinegar based on the criteria that it had a cost of 0.649 moles per dollar compared to the cost of the gourmet vinegar, 0.207 moles per dollar. Secondly, the concentrations of sodium bicarbonate and sodium carbonate were determined within the unknown as 0.5507 M and 0.246 M, respectively.

Pre-Lab Questions

1) You want to prepare 50 mL of 0.10 M NaOH from exactly 50 percent (wt/wt) NaOH solution. How many milliliters of 50 percent solution are required?

   Total moles needed in desired solution: $0.050 \text{ L} \times \frac{0.10 \text{ mol}}{\text{L}} = 0.0050 \text{ mol}$

Weight: $0.0050 \text{ mol} \times 36.997 \text{ g/mol NaOH} = 0.20 \text{ g NaOH}$

50% solution of NaOH → $\frac{0.20 \text{ g NaOH}}{0.40 \text{ g solution}}$ Therefore, 0.4 grams of solution is needed.
Density of 50% NaOH at 25 °C: 1.52170 g/mL

$$\frac{0.40 \text{ g}}{1.52170 \frac{g}{mL}} = \boxed{0.26 \text{ mL of 50\% NaOH}}$$

2) What is a primary standard? Can KOH be used as a primary standard? Explain.

A primary standard is the solid chemical that is used to standardize the secondary standard, the titrant solution. This solid is weighed, dissolved, and brought to a known volume, and through its titration, calculations can be conducted and the secondary standard can be accurately standardized. Primary standards work best when they have a high molecular weight, do not absorb molecules or react with the air, and can be stored at high levels of purity. Under these criteria, KOH would not be a good primary standard; its molecular weight is only 56 g/mol, and it reacts with and absorbs CO_2 readily from the air. When titrating, it is necessary that the only reactions occurring are those between the titrant and analyte; if KOH was used, other reactions would be occurring.

3) An aqueous sample is analyzed and found to contain 4 percent NaOH (4 g in 96 mL H_2O)

a) Calculate the concentration of the solution in molarity (assume the density of water is 1.00 g/mL and the volume of the solution is 100 mL)

$$\frac{4 \text{ g NaOH}}{39.997 \text{ g/mol}} = 0.100 \text{ mol NaOH}$$

$$\frac{4 \text{ g NaOH}}{0.1 \text{ L solution}} = \frac{0.100 \text{ mol}}{0.1 \text{ L}} = \boxed{1.00 \text{ M concentration}}$$

b) If this solution is neutralized with 1 M sulfuric acid, how many mL of sulfuric acid will be needed?

H_2SO_4 produces 2 moles of H^+ ions for every 1 mole of H_2SO_4 present; there is 1 mol of OH produced for every 1 mole of NaOH. Since the molar concentrations are the same, 100 mL of 1 M sulfuric acid has the same number of moles of solute as 100 mL of 1 M NaOH. Thus, the amount of sulfuric acid needed to neutralize the solution is half of the amount of NaOH present:

$\boxed{50 \text{ mL needed to neutralize 100 mL of 1 M NaOH}}$

c) What is the normality of the NaOH solution?

Normality = molarity × equivalents (the number of moles of reactive atoms/molecules) → 1 M NaOH; 1 mol OH (reactive molecule) produced.

Normality = 1 × 1 = $\boxed{\text{1 N}}$

Post-Lab Questions:

1. What precautions did you take in setting up a micro or a macro buret?

In this lab, the macro buret was carefully attached to the buret stand so that it would not fall out during the titration. One of the most important safety precautions was taking the buret out of the clamp and using a funnel when pouring strong acid or base into the buret. The buret itself was also held while operating the stopcock during the titration to ensure that the buret would not topple out of the clamp.

2. What amount of oxalic acid ($H_2C_2O_4 \bullet 2H_2O$, MW = 126.0 g/mol) must be used to prepare 10 mL of 0.130 M oxalic acid solution? This acid has two replaceable acidic protons. What would be the normality of the acid?

$\frac{0.130 \text{ mol}}{1000 \text{ mL}}$ × 10 mL = 0.0013 mol oxalic acid needed

0.0013 mol × 126.0 g/mol = $\boxed{\text{0.16 g oxalic acid}}$

0.130 M × 2 = $\boxed{\text{0.26 N}}$

Since oxalic acid can donate two protons, the molarity is multiplied by a factor of 2 to obtain the normality of this solution.

3. In you take 50 mL of 5 percent acetic acid, how many millimoles of acetic acid have you taken?

50 mL solution × 1.005 g/mL = 50.25 g of solution

0.050 × 50.25 g solution = 2.5125 g acetic acid

$\frac{2.5125 \text{ g acetic acid}}{60.05 \text{ g/mol}}$ × $\frac{10^3 \text{ mmol}}{1.000 \text{ mol}}$ = 41.84 mmol = $\boxed{\text{42 mmol}}$