Determining K_a for Ethanoic (Acetic) acid

Abstract

A titration lab is a technique that is used to determine the concentration of an unknown solution by using a solution with a known concentration (15). Titration methods have been used since the beginning of the 18th century for volumetric analysis as the study began in France and Francois Descroizilles was the first person to make and use a burette (2). Titration experiments are very delicate labs as very small amounts of differences in volume or misreadings can affect the results and alter calculations.

This lab will contains phenolphthalein and it is a very popular chemical used during titration labs to identify the level of pH and to identify when the endpoint has reached (5). Universal indicators strips react with OH^- or H+ ions, and depending on the amount of these ions that are present, a different colour would be produced associated with the pH level (5). Phenolphthalein was added inside of the ethanoic acid solution to indicate when the mixture had turned basic and end point was reached. This pink colour occurs because the phenolphthalein is a weak acid itself, and so when it loses H+ ions because of the base, the ions of the phenolphthalein are pink (3).

In this experiment, the acid ionization constant (K_a) of acetic acid was calculated through the results obtained from the titration experiment. The solution in the burette was the sodium hydroxide, which is a base and it was added to 10 mL of acetic acid until it reached the end point. The end point of a titration lab is when the solution being tested on, the analyte (acetic acid) changes its colour (14). The equivalence point of a titration lab occurs right before, which means that the two solutions in the container, both the analyte and the titrant are chemically equal in moles, and therefore the solution is neutralized (4) .

The results from the titration that showed the amount of volume needed to reach equivalence point can be used to determine the acid ionization constant in this experiment. The K_a of an acid is a value that is used to predict the acid dissociation (7). This means that a larger K_a value indicates that the acid is stronger because more of the acid dissociates, while a lower K_a value means that the acid is weak, because the acid is not dissociating as much. In strong acids, complete dissociation occurs (7) .

If acetic acid was added in a solution with water, the water would act as the base, because the acetic acid will be the proton donor, while the water will be the proton acceptor (8). A hydronium ion will be produced, which is a protonated water molecule that is present in all the aqueous acids (9). Using the concentration of the hydronium ion and the pH of the acid, the K_a value of the acid can be calculated.

Both the reactants that will be used in the titration experiment are very known compounds. Ethanoic acid, also known as acetic acid is a weak acid that is the main ingredient in vinegar that gives it the define taste and smell (10). The compound is also used to make perfumes, dyes and inks. Sodium hydroxide will be the other reactant used in the experiment which is a strong base (13). This compound is used in chemical cleaners and in the manufacturing of plastic, paper and soap (13).

There were a few safety precautions taken in this experiment when handing either of the two compounds. Ethanoic acid is a corrosive and flammable substance (1). If contact is made with skin, wash

skin thoroughly. If contact is made with eyes, rinse eyes for at least 15 minutes. If inhaled, move to fresh air and loosen clothing, and if ingested do not induce vomiting. Seek medical attention if needed (1). Sodium hydroxide is a corrosive compound that can cause severe skin burns and damage if contact is made. Handling of sodium hydroxide is similar to handling of acetic acid, and take same safety measures. Sodium hydroxide is damaging to aquatic environments, and should be disposed of properly (11).

Purpose

The purpose of this lab was to determine the acidity constant (K_a value) of acetic acid by the use of a titration experiment, by adding the acetic acid to a solution of sodium hydroxide.

Hypothesis

If a certain volume of sodium hydroxide (NaOH 0.25M) is added to 10 mL acetic acid in a titration experiment, than the calculated K_a value will be 1.8×10^{-5} (14) .

Materials and Apparatus

Materials used

- 6 drops of Phenolphthalein
- 25 mL of sodium hydroxide solution (0.25M)
- 10 mL of acetic acid solution (0.25M)

Apparatus used

- 1 125 mL erlenmeyer flask
- 1 Burette
- 1 Burette clamp
- 1 Retort stand
- 2 pH universal indicator papers
- 1 10 mL pipette
- 1 funnel
- Safety goggles

Procedure

Refer to the pages 552 and 553 for the complete procedure from the Mcgraw-Hill Ryerson Chemistry 12 textbook.

Observations

Table 1: Qualitative Observations of physical properties and changes of the reactants and products used before, during and after titration.

Solution	*Before Titration*	*During Titration*	*After Titration*	*pH colour*
NaOH (sodium hydroxide)	*Colourless Transparent Slightly more viscous than water*	*Colourless Transparent*	*N/A*	*N/A*
CH_3COOH (ethanoic acid)	*Transparent Colourless Vinegar odour*	*The solution remained the same, there was no significant change from how it appeared before titration. At certain moments solution turned into a faint pink, than returned to being transparent.*	*N/A*	*When the pH indicator was dapped into the solution, it had turned to a red/orange colour, indicating it was about 3.0*
CH_3COONa (*Sodium acetate)*	*N/A*	*N/A*	*After the two reactants were mixed, the solution changed to a transparent faint pink colour and remained that way. This occured in both trials.*	*N/A*

Table 2 : Quantitative observations of the volume and pH of NaOH and CH_3COOH, before and after the titration.

Trial #	*Initial volume of NaOH (mL)*	*Final Volume of NaOH (mL)*	*Volume of NaOH used (mL)*	*pH level of acetic acid according to universal indicator*	*Volume used of acetic acid (mL)*
1	*6.0*	*17.5*	*11.5*	*3.0*	*10.0*
2	*17.5*	*28.7*	*11.2*	*3.0*	*10.0*

Calculations

Balanced equation

$CH_3COOH_{(aq)} + NaOH_{(aq)} \rightleftharpoons CH_3COONa_{(aq)} + H_2O_{(aq)}$

Average volume of NaOH used

V= 11.50mL + 11.20 mL / 2

V= 11.35mL

Moles of sodium hydroxide

n= C x V

n=(0.25M)(0.01135L)

$n=2.8 \times 10^{-3}$ *mol*

The molar ratio of sodium hydroxide to the acetic acid is 1:1, so this value for moles can be used for the calculations for the concentration of acetic acid.

Concentration of Acetic Acid

C= n/v

$C= 2.8 \times 10^{-3}$ *mol/ 0.010 L*

C= 0.28 M

Concentration of hydronium ion

$[H_3O^+] = 10^{-ph}$

$= 10^{-3}$

$= 1.0 \times 10^{-3} M$

Dissociation equation of Acetic acid

$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$

Ice Chart	CH_3COOH	H_3O^+	CH_3COO^-
I	*0.28M*	*0*	*0*
C	*-x*	*x*	*x*
E	*0.28-x*	*x*	*x*

Calculating K_a of acetic acid

K_a *= [Products] / [Reactants]*

$=[H_3O^+][CH_3COO^-] / [CH_3COOH]$
$=[x^2]/ [0.28-x]$
$=[1.0 \times 10^{-3} M]^2/ [0.28M- 1.0 \times 10^{-3}M]$
$K_a= 3.6 \times 10^{-6}$

Therefore the calculated K_a value using the volume and pH obtained from the titration lab is 3.6×10^{-6}.

% ionized
= [[0.0010M]/ [0.28M]] x 100%
= 0.36%

<u>Percent error of the K_a value of acetic acid</u>
% error calculation = [experimental - expected / expected] x 100%
$= [3.6 \times 10^{-6} - 1.8 \times 10^{-5} / 1.8 \times 10^{-5}] \times 100\%$
$= 8.0 \times 10^{1}$ %

<u>Calculations of what the pH value should have been using the theoretical K_a value</u>
Moles of sodium hydroxide
n= C x V
*n=(0.25M)(**<u>0.010</u>**L)*
$n=2.5 \times 10^{-3}$ *mol*

Concentration of Acetic Acid
C= n/v
$C= 2.5 \times 10^{-3}$ *mol/ 0.010 L*
C= 0.25 M
The concentration of the acetic acid was 0.25M as indicated later on the flask. This concentration is achieved mathematically if **<u>10</u>** mL of the sodium hydroxide was added, instead of 11.35mL.

Acetic acid dissociation equation

$$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$$

Ice Chart	CH_3COOH	H_3O^+	CH_3COO^-
I	*0.25M*	*0*	*0*
C	*-x*	*x*	*x*
E	*0.25M*	*x*	*x*

The theoretical K_a value of the acetic acid is 1.8×10^{-5}.

$K_a= [H_3O^+][CH_3COO^-] / [CH_3COOH]$

$1.8 \times 10^{-5} = x^2 / 0.25M$
$X^2 = (0.25M)(1.8 \times 10^{-5})$
$X = [H_3O^+] = 2.12 \times 10^{-3}$ M

pH= $-\log[2.12 \times 10^{-3}M]$
pH= 2.67

Therefore the pH level was about 2.67 for the acetic solution that had a concentration of 0.25M theoretically.

Discussion

The performed lab was unsuccessful as the correct acid ionization constant for ethanoic acid was not calculated. It was hypothesized that the ethanoic acid would have a K_a value of 1.8×10^{-5} , but the results from the calculations showed that the K_a value was 3.6×10^{-6}. Furthermore, the calculated percent error showed that there was 80% error, which means there must have been error that occured in the lab. For these reasons the experiment was not a success.

To identify the main sources of error that may have taken place during the titration, it is important to note the variables that will affect the final calculation of the $K_{a,}$ value. Firsty, the pH value identified of the ethanoic acid was inaccurate, which caused the K_a value to be less than it should had been. The pH indicator used was a universal indicator, which allows someone to have a general idea of what the pH is, as it is not very accurate (no decimals). This makes the reading of of the universal indicator dependant on the perception of the person reading it. The theoretical value of the K_a was 1.8×10^{-5} and the calculated K_a was 3.6×10^{-6}. This value was obtained by the use of previous calculations that involved the pH that was used to determine the hydronium ion concentration. The pH value was 3, but if the pH value was read to be 2.67 than the calculated value of the K_a would be 1.8×10^{-5}, which is the exact value of the theoretical K_a value of the acetic acid. To see how the difference in pH affected the calculations for K_a, refer to the calculations section for the complete solution. If the pH was indeed read properly, than the percent error would have been 0%, which is a very accurate value. This error shows how delicate this lab is, because a tiny change of 0.35 in pH value changed the K_a value by a large amount.

The second variable that may have affected the calculated K_a value is the volume of the sodium hydroxide used, because it was part of the calculation to find molarity for the reactants, which is than used to calculate the K_a value. The way the sodium hydroxide was being added into the ethanoic acid was by opening the valve from the burette, allowing drops to enter the analyte and swirling the erlenmeyer flask. An error may have occurred because too much base was added to the acid, more than when the endpoint was reached. This could have occured because of two reasons. First, the erlenmeyer flask was not being mixed thoroughly throughout the entire time drops were being added and so it was not seen when the mixture should have changed colour. Second, the base was not being added in small proportions to the flask, which means there may have been a few more mL added than needed after reaching the end point. The average of the volume of sodium hydroxide needed to reach the end point of the titration was about 11.35mL when actually it should have been 10 mL. The concentration of the acetic acid was revealed after the experiment, and it was 0.25M, same as the concentration of the sodium hydroxide. Since the two reactants had the same molar ratio, the volume used for the titration should have been the same as well, and 10 mL were used for the acetic acid. This means that if 10 mL was used to calculate the number of moles in sodium hydroxide, than the correct value for the concentration of acetic acid would have been achieved, which is 0.25M, instead of 0.28M. Combining the actual volume that should have been used with the actual pH level calculated, and the exact result for the K_a value of the acetic acid is achieved. To see full solution as to how the volume affected the concentration of the acetic acid, refer to the calculations section.

The third source of error is that the theoretical K_a value of the acetic acid that the calculated one is being compared to is at SATP. This refers to a temperature of 25℃ which is the variable that can affect the K_a value (12). These were not the exact conditions of the classroom and therefore this will alter the theoretical K_a value of the acetic acid. The concentration used to calculate the K_a was obtained from the the calculation of the concentration of hydronium ions, which used pH. The pH of a substance is affected by temperature, which means at different temperatures the pH can be higher or lower of the same substance and the K_a value will also be different (6). According to westlab, when temperature increases molecular vibrations will also increase (6). This allows the water produce a greater number hydrogen ions which will result in a decreased pH level, because the more hydrogen ions released, the more acidic a substance is (6). In this lab, the theoretical value of K_a was calculated in a classroom that is roughly 21℃, which is lower than 25℃. The pH obtained in the classroom is greater (3.0) than the theoretical value which had a pH of 2.67 because the classroom is colder than SATP, meaning hydrogen ions are not released as much which makes the pH greater. As mentioned earlier, the K_a is dependant on the pH level because each K_a value of a substance is associated with a specific pH, and that is associated with a specific temperature. This means if the theoretical value being compared to the calculated value was obtained under the same conditions of the classroom than the two results would most likely be more similar.

For future labs, there are a few ways to ensure results are more reliable and precise. The first suggestion is to use a pH meter instead of the pH universal indicators to determine what the pH level of the acid was. Firstly, the pH meter is not dependant on someone's vision and perception of colours because the meter gives the pH number level directly, instead of having to associate the colour change with the level of pH. Secondly the pH meter also reads to a decimal place, which provides even more accuracy to what the pH level actually is. This suggestion eliminates any sources of error associated with identifying the pH level of the acetic acid. Another suggestion is to take more time when performing the titration, and go slower when dropping the base into the acid from the burette, in very small amounts. Have someone continually swirling the erlenmeyer flask while another person is ready to close the valve, instead of one person managing both tasks. This allows the experimenters to see the exact moment the end point was reached without overshooting, providing a more accurate volume. Lastly conducting the experiment at the same temperature that the theoretical $K_{a,}$ calculation was conducted in (25℃) to make the titration lab more fair and precise.

Conclusion

In conclusion, the hypothesis was proven to not be correct. During the titration, the two most important variables were the pH and the volume used to achieve neutralization between 10 mL of acetic acid and sodium hydroxide (0.25M). The calculated K_a value was 3.6×10^{-6}, although it should have been closer to 1.8×10^{-5}. This lab proved to have numerous errors that could have been the reason such a result was achieved, and a percent error of 80%. The pH level was misread which affected the calculation for concentration, the overshot amount of volume of sodium hydroxide needed to neutralize the acetic acid and the proper conditions such as SATP temperature was not met, were all reason why the K_a value was less than the theoretical result.

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