Synthesis of salicylic acid and potentiometric determination of its purity and di...

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Abstract

The purpose of the study is to synthesize salicylic acid from the ester, methyl salicylate, and determine the acid's dissociation constant and purity. The ester was converted to salicylic acid by base hydrolysis. The products were refluxed and recrystallized, to ensure maximum purity, and filtered, dried, and weighed. The melting point of the product was determined using a Fischer-Johns melting point apparatus.

The acid then dissolved in separate beakers with 95% ethanol and water and titrated with 0. 050 M NaOH, previously standardized with potassium hydrogen phthalate, through potentiometric titration. The pH after addition of base was measured and plotted against the volume of titrant added using three different plots. Results show a 61. 0% yield and the melting point differed from the theoretical by a range of 3. 11-6. 83%. The pKa calculated was 2. 865, differing from the literature value of 2. 98, by 3. 86%. The theoretical purity of the sample was 100. 0%, which differed with the experimental one by 1. %; the experimental purity is 101. 7%. Potentiometric titration proves to be adequate in the determination of the acid dissociation constant and purity of a sample. Aside from that, the synthesis proved adequate given the high purity of the product. Keywords: acid dissociation constant purity melting point ester

INTRODUCTION

Potentiometric methods of analysis are based on measuring the potential of electrochemical cells without drawing much, appreciable current. For centuries, potentiometry has been used to locate the endpoint in most

titration set-ups. Skoog, et al. , 2004). Potentiometric methods offer a myriad of advantages, its main advantage being its low operational costs. Voltmeters and electrodes are, generally, far cheaper than most modern scientific instruments. Models suitable for direct potentiometry in field work, away from the laboratory, are inexpensive, compact, and easy to use. Essentially nondestructive of the sample, i. e. , insertion of the electrode does not drastically change the composition of the test solution (except for the slight leakage of electrolytes from the reference electrode), electrodes are relatively free from interferences.

Stable potential readings are attained fairly rapidly and voltages are easily recorded as functions of time. Finally, the wide range of analyte activities, over which some of the available indicator electrodes exhibit stable, nearly Nernstian responses, represents an important advantage (Day and Underwood, 1991). Potentiometric titrations involve measurement of the potential of a suitable indicator electrode as a function of titrant volume. This indicator electrode is speecific to the hydronium ion, H3O+.

It offers adavantages over direct potentiometry in that it is not dependent on measuring absolute values of Ecell. This is due to having the measurement based on the titrant volume that causes a rapid change in potential near the equivalence point. This makes the process relatively free from the juction potential uncertainties as this potential remains approximately constant during the titration process. Instead, the titration results depend heavily on having a titrant with accurately known concentration. The instrument merely signals the endpoint and behaves like a regular chemical indicator.

Aside from that, the reference electrode potential need not be known. Most importantly, ionic strength effects are not important in the titration procedure because the result is analyte concentration, even if the electrode responds to activities. The dissociation of a weak monoprotic acid is given by the equation HA - H+ + A– (1) [H+][A–] [HA] where HA is the monoprotic acid, and A– is its conjugate base. The corresponding equilibrium constant for this acid dissociation is defined as Ka = (2)

Ka[HA] [A–] which, when seeking to find [H+], rearranges to [H+] = (3) Prior to the equivalence point, before any titrant was added and the analyte is the only species present in solution, the pH of the system is calculated from the concentration of that analyte and its dissociation constant. In the case where a weak monoprotic acid is being titrated with a strong base, subsequent addition of the titrant will cause a reaction to occur between the acid and the base. HA + OH– > H2O + A– (4)

The presence of the acid and its conjugate base in solution will cause the formation of a buffer solution, which are solutions that resist a drastic change in pH, should a strong acid or base be added to the system. At these points, the slope of a constructed titration curve is at its minimum. This is the pre-equivalence point. There is a point during the pre-equivalence point region wherein both the acid and its conjugate base are present in equal amounts. This occurs when half of the acid has been neutralized, or when the titration is at the half-equivalence point.

At this point, the buffering capacity of the system is at its maximum. Aside from that, simplifying Eq. 3 at this point by inspection, the [H+] is equal to that of the Ka. Taking the negative of the logarithm of both sides, one will be able to get the relation pH = pKa (Skoog, et al. , 2004). Thus, the dissociation constant may be computed through determining the pH at half-equivalence point. This experiment will use salicylic acid as the analyte. Salicylic acid, Fig. 1. Salicylic acid is a weak monoprotic acid, capable of releasing the acidic hydrogen connected to the –COOH group.

The hydrogen of the phenol group is hard to release because the oxygen is stabilized by resonance. Salicylic acid is a naturally occuring substance, usually produced by plants. It is found mainly in the willow's leaves and bark. The pure acid possesses several useful medicinal properties. It is an antipyretic (fever reducer), analgesic (pain reliever) and anti-inflammatory (swelling reducer). However, pure salicylic acid makes for an extremely unpleasant medicine. Salicylic acid contains two acidic functional groups, the phenolic (C6H5OH) and the carboxylic acid (RCOOH) groups.

These groups cause the acid to be an irritating substance that burns the delicate lining of the mouth, throat, and stomach, hence its esterification to acetylsalicylic acid or aspirin, before ingestion as analgesic (Reed College, 2009). Esterificaton is the process by which a carboxylic acid is transformed to an ester. Esters are organic compounds that are derived usually by reacting a carboxylic acid and an alcohol. The general formula for esters is with the RC= O group derived from the parent carboxylic acid, and the –OR’ group from the parent alcohol.

The mechanism for the reaction of the alcohol and carboxylic acid to form the ester is as follows: Fig. 2. Mechanism of Esterification from a Carboxylic acid RCOOH and alcohol R’OH Esters may also be synthesized by reacting the carboxylic acid with other reagents such as SOCl2 to form the acyl chloride, which will then be treated with an alcohol in pyridine, to esterify it. Esters are among the most widespread of all naturally occuring compounds. Many esters are pleasant-smelling liquids that are responsible for the fragrant odor of fruits and flowers. For example, methyl butanoate is found n pineapple oil and isopentyl acetate is a constituent of banana oil. The ester linkage is also present in animal fats and other biologically important molecules. The chemical industry also uses esters for a variety of purposes. Ethyl acetate, for example, is commonly used as a solvent while many dialkyl phthalates are used as plasticizers to keep polymers from being brittle (McMurry, 2004). Methyl salicylate is produced by many plants. It was first isolated from wintergreen leaves, Gaulthea procumbens, and is commonly known as oil of wintergreen. Fig. 3. Methyl salicylate

An ester of salicylic acid and methanol, it masks one of the acidic hydrogens in salicylic acid by replacing it with a methyl (CH3–) group. Hence, it is a relatively unreactive compound that does not release salicylic acid efficiently into the body. It is, therefore, not an effective analgesic, or pain-killer. However, it is added to many products, notably for its fragrance, especially root beer and liniments. In order for it to be activated, methyl salicylate must be converted to salicylic acid by organic synthesis, specifically through

saponification, a process not unlike that undergone by animal fats to become soaps (McMurry, 2004).

In this experiment, salicylic acid will be synthesized from methyl salicylate by base hydrolysis. Its dissociation constant will also be measured through potentiometric titration. Likewise, the percentage of purity the salicylic acid used in the reaction will be likewise determined.

METHODOLOGY

In synthesizing salicylic acid, 1. 2 g of sodium hydroxide (NaOH) were dissolved in 7 mL water in a round bottom flask.. Half of a milliliter (0. 5 mL) of methyl salicylate was added to this mixture. The mixture was then efluxed for 15 minutes and cooled to room temperature. One-milliliter increments of 3 M sulfuric acid (H2SO4) were added until the formation of a white precipitate, salicylic acid. Half of a milliliter (0. 5 mL) of the acid was added to ensure complete precipitation of the product. The mixture was then cooled in an ice water bath with a temperature of at most 5°C for the reaction to subside. The product was then filtered and rinsed with cold water, and recrystallized in water. The solids were then filtered on a pre-weighed filter paper and air-dried in the locker.

When dried, the solids, along with the filter paper, were weighed and the melting point determined. Two hundred and fifty milliliters (250 mL) of 0. 50 M stock sodium hydroxide (NaOH) solution was then prepared. From this solution, 250 mL of 0. 050 M standard solution was prepared for the titration. This solution was then standardized with 0. 2000 g 99. 95% potassium hydrogen phthalate (KHP) and phenolphthalein as indicator to the pale pink

endpoint. Two more trials were done. Three samples of pure salicylic acid were weighed at approximately 0. 000 g and placed in three separate 250-mL beakers. A minimum amount of 95% ethanol was then added to dissolve the acid, and was diluted with 50. 00 mL distilled water. The electrode was first calibrated such that it would relate the developed potential to the pH. At this point, the potential would be measured as pH when increments of the titrant are successively added. The base burette, pH meter, and magnetic stirrer were set up according to Fig. 6. Fig. 6. Potentiometric titration set-up A spin bar was place inside the beaker with the sample solution.

The electrode was also positioned such that it would not get hit with the spin bar. The pH of the set-up was first measured prior to addition of base. For the first trial, 1. 00 mL of the titrant was added to the analyte and the pH recorded. This was done until the pH was 11. 50. From there, the equivalence point was approximated based on the volume of the titrant that caused a large change in pH. For the two succeeding trials, 1. 00 mL of the titrant was added to the analyte. At ±5. 00 mL of the equivalence point, the volume of titrant added was reduced to 0. 0 mL increments. At ±3. 00 mL and ±2. 00 mL of the equivalence point, it was further reduced to 0. 20 mL and 0. 10 mL, respectively. The titration contnued beyond 5. 00 mL of the equivalence point using 0. 50 mL of titrant until the pH registered was 11. 00.

RESULTS AND DISCUSSION

Methyl salicylate reacts with a strong base in the following manner: Fig. 4. Base hydrolysis of methyl salicylate The methoxy group of the ester is

substituted by the hydroxide ion through nucleophilic acyl substitution. The hydroxide attaches itself to the carboxylic carbon.

A fast proton transfer from the hydroxide to the methoxy group occurs such that the methoxy gets protonated and leaves the substrate. The sodium ions stabilize the negative charge of the salicylate ion predominantly found near the two oxygens of the ion. The solution was then refluxed to prevent loss of material and to prevent the inclusion of impurities in the product. After cooling to room temperature, 1. 00 mL of 3M sulfuric acid (H2SO4) until salicylic acid starts precipitating as a white solid. To complete the precipitation, 0. 50 mL of the strong acid was added to the mixture.

At this point, the salicylate ion is protonated and the final product, salicylic acid, forms through the reaction: Fig. 5. Protonation of the phenolate and carboxylate groups of the salicylate ion The flask containing the precipitate was then doused in cold water to stop the reaction. Cold water was used in rinsing the solid after filtration to wash out impurities that were insoluble in the solvent. The solid was recrystallized in hot water. Dissolving the solids in hot water generally increases the solubility of the substances, hence the solids dissolve along with the soluble impurities.

The mixture was then allowed to cool slowly. As the solution cools, the solubility of the compound (and of the soluble impurities) decreases, the solution becomes saturated with the desired compound, and the compound begins to crystallize. Because formation of crystals is a highly selective process that usually excludes foreign molecules, only crystals of the desired

compound form. Because the soluble impurities are present in smaller amounts, the solution never becomes saturated with the impurities, so the impurities remain in solution even after the solution has cooled.

Removing the solution from the crystals thus removes the solvent and the soluble impurities from the desired crystals. A final rinse of minimum ice water ensures the displacement of the impurities still clinging on the surface of the product (“ Recrystallization,” accessed 6 Sept 2010). After it was recystallized, the solid was filtered again in a pre-weighed filter paper. The filter paper used in filtering the solid weighed 1. 0349 ± 0. 0002 g. The mass of the dried product and the filter paper was 1. 3610 ± 0. 0002 g; this means that 0. 3261 ± 0. 0003 g was synthesized.

Theoretically, with methyl salicylate as the limiting reagent, the mass of salicylic acid should be 0. 5400 g. Computing for the %yield, would give 61. 0%. In the determination of the melting point, the observed range of the melting point was 150-156°C. Comparing with the theoretical value, 161°C, the relative error lies within the range 3. 11-6. 83%. Therefore, the synthesized salicylic acid has a high purity as exhibited by the magnitude of the relative error. This could justify the %yield given that recrystallization might have caused a loss of material due to the increase of the number of steps involved.

It should be noted that in organic synthesis, steps leading to the target molecule do not give 100% yield, hence, increasing the number of steps would decrease the yield. Aside from that, the low yield could be attributed

to the accidental rapid cooling that occurred during the recrystallization process. In a titration curve, there are three important regions: pre-equivalence point, equivalence point, and post-equivalence point. For the titration of salicylic acid with NaOH, the pre-equivalence point is characterized by the dependence of pH on the concentration of salicylic acid.

Let salicylic acid be HA; prior to addition of the base, the pH can be calculated by the concentration of HA and its acid dissociation constant, Ka. When the base is added, HA reacts with the hydroxide ion (OH–) according to the Eq. 4 stated above. Therefore, one has a buffer solution comprised of salicylic acid and salicylate. At equivalence point, all of the acid has been converted into its conjugate base. This base will hydrolyze with water through the reaction A– + H2O > HA + OH– (6) reforming a small portion of the acid. At the equivalence point, the slope of the titration curve is at its steepest.

The pH at this point is dependent upon the conjugate base formed and its base dissociation constant, Kb. Beyond the equivalence point, the post-equivalence point, the pH is dependent on the concentration of the excess titrant added. This is because the contribution of the conjugate base is very small and, therefore, negligible compared to that of sodium hydroxide. However, salicylate is still the dominant species of salicylic acid present in solution. A distinction between equivalence point and endpoint should be established when titrating.

Skoog, (2004), states that the equivalence point is the point in titration where the amount of titrant added is chemically equivalent to that of the analyte in solution. The endpoint is a point during titration wherein an observable change signals that the amount of titrant added is chemically equivalent with that of the analyte. The endpoint may not necessarily coincide with the equivalence point, especially in neutralization titrations. It may come earlier or later than the equivalence point, depending on the indicator used. This difference pf volume at equivalent and endpoint is known as titration error.

In potentiometric titration, no chemical indicator is required. Instead, the endpoint is characterized by the drastic change in pX, measured by the electrode. X may be H3O+, OH–, a cation or anion, or any substance participating in the titrimetric reaction (Day and Underwood, 1991). In this case, the volume of titrant that contributes to the greatest change in pH is the volume required to completely titrate the salicylic acid in solution. Thus, the increments of addition of titrant are narrowed down as the equivalence point is reached because that way, the equivalence point will not be missed.

Aside from that, the exact volume of titrant required for complete reaction of the analyte would be detected. Prior to titration and dilution with water, the salicylic acid was dissolved with a minimal amount of 95% ethanol. Salicylic acid is sparingly soluble in water, a polar solvent that can hydrogen-bond with solutes that can hydrogen-bond with its molecules. This is due to the acid having more than 6 carbons increasing its non-polarity, although it has

oxygen that can form hydrogen bonds with water (McMurry, 2004). Adding ethanol reduces the polarity of the solvent, facilitating dissolution of the acid.

It was noted that some of the acid reprecipitated upon addition of water. Thus, a minimum amount of 95% ethanol was again added to dissolve the acid. In the experiment, the electrode was also not lowered while the bar was spinning as air bubbles would adhere on its surface causing an error in the measurement of pH. Constant stirring is important in potentiometric titrations, as in other types of titrimetric analyses. Stirring is done because it will cause the titrant to react not just with the surface of the analyte where it dropped but with the entirety of the analyte solution.

This way, the reaction will go to completion and minimal error will be observed. The volumes used for standardization of NaOH with KHP as primary standard are tabulated in Table 2 in the Appendix. Standardization of sodium hydroxide gave 0. 04633 M NaOH. This value was used in the determination of the percentage of purity (%purity) of the salicylic acid sample. For the determination of the equivalence point volume, three plots were done for each trial. The first plot is the graph between pH vs. volume of titrant. The second is the first derivative plot with which ? pH/?

V was plotted against V’, where ? pH and ? V are the change in pH and the volume added between two consecutive readings, respectively, and V’ is the average volume between two consecutive readings. The third plot is the second derivative plot between ? 2pH/? V2 and V”. The ordinate of the graph, ? 2pH/? V2, was obtained by taking the quotient of the difference

between two consecutive readings of ? pH/? V, and ? V, whereas V” is the average of the two consecutive readings of V’. Only two trials were done because of the inadequacy of the time. The graphs for the first trial are: a) (b) (c) Fig. 7. Titration Curves for the First Trial. (a) pH vs Vtitrant; (b) ? pH/? V vs V’; (c) ? 2pH/? V2 vs V” MNaOHVNaOH at eq ptFWsalicylic acid Gsample The graph for the second trial may be found at the Appendix. Either of the three graphs of Figs. 7 and 8 may be used in the determination of %purity for each trial. The equation used for determining the %purity, specifically for this experiment is %purity = x 100 (7) Thus, the volume of titrant used at equivalence point is required for the calculation.

The interpolated values of the volume at equivalence point are tabulated in Table 3 in the Appendix. Calculation of the mean %purity is 101. 7%, having an error of 1. 7%, as the theoretical value is 100. 0%. The pKa is based on the pH at half-equivalence point by virtue of the derivation of Eq. 3. Onle Figs. 7a and 8a may be used for the determination of pKa as the other four do not directly give the pH at each point. The interpolated values for the pH at half-equivalence point is tabulated at Table 3 in the Appendix. The experimental pKa is 2. 865. Therefore, the Ka of the acid, given Ka = – log[Ka] (7) or Ka = 10–pKa (8) is 1. 3646 x 10–3. The theoretical pKa is equal to 2. 98. Thus, the pKa value gave rise to a 3. 86% error. Statistical analysis of the results shows the following results: Table 1. Statistical Analysis of the Results | Range| Standard Deviation| %purity| 15. 7%| 11. 1%| pKa| 0. 21| 0. 148| | Relative Standard Deviation (RSD), ppt| Confidence Intervals (95% confidence)| %purity| 109. 1| 101. 7 ± 99. 7 %| pKa| 51. 8| 2. 865 ± 1. 334|

The RSD of %purity is relatively large such that the values have low precision.

Aside from that, the confidence intervals for the mean is also large such that it almost has the same order of magnitude as that of the mean. This means that the mean exhibits very low accuracy. With regards to the pKa, the RSD showed a low value, only 51. 8 ppt, implying high precision amongst the values. On the other hand, the mean value shows low accuracy because of the magnitude of the confidence intervals. These errors could be attributed to the number of trials. Due to inadequate time, the group was only able to do 2 trials, one less than the prescribed number of trials.

CONCLUSIONS

The experiment aimed at synthesizing salicylic acid from methyl salicylate and determining the acid dissociation constant Ka of the acid along with its purity. Organic synthesis provided a 61% yield of the acid, a relatively low yield. However, the purity of the acid can be classified as high due to the observed melting point range’s precision with the theoretical one; the %differenceis only 3. 11-6. 38%. The results of the potentiometric titration show that the salicylic acid used was 101. 7% pure, a value greater than the purity, which is 100%.

The Ka, expressed as pKa, obtained was 2. 865, 0. 035 units less than the theoretical value, which is 2. 98. Though the %differences are low, the accuracy of the computed values is questionable given that the confidence intervals for the %purity and pKa are ± 99. 7 and ± 1. 334, respectively. These errors can be attributed to the number of trials done, which is two,

due to the slow stabilization of the pH meter readings that resulted in a long period of titration. Thus, it can be concluded that potentiometric titration is an effective way of determining the acid dissociation constant of a sample.

It is recommended that a better pH meter be used in the measurement of the pH and that solutions used be titrated immediately.

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