Experiment on acid-based reactions in aqueous solution

LAB S. ACIDS AND BASES: TITRATIONS

Introduction:

Within the divisions of chemistry there are several distinct types. The subject of study within this report are acid-base reactions in aqueous solution. Antoine Lavoisier, a French chemist, was the first to study acid base reactions within 1776 [1] . More recently, the conceptualization of acids and bases has been refined by scientists Arrhenius and Bronsted Lowry, who defined the idea of what an acid and base were. According to the Bronsted-Lowry definition of acids and bases, an acid is a substance that " donates" a proton (or an H^+ ion) to water when in solution, and a base is a substance that " borrows" an H^+ ion when in solution [2] . When an acid donates this proton, it bonds with a water molecule and creates an H_3O^+ , or Hydronium ion, and when a base borrows a proton, the water molecule changes to a Hydroxide ion, OH^- [1] . When an acid and a base are mixed a neutralization, process occurs. Within a neutralization reaction, it is a paramount that the moles of the acid are equal to the moles of the corresponding base.

Acids all have a conjugate base, which is what they react to in water, and they are either strong or weak. Conjugate acid-base pairs are related to one another through the gain or loss of one proton. Every acid-base reaction will have two acid-base pairs. Some acids are better at donating than others. If they donate really well (100% of their protons,) then they are called strong acids. Whereas, when bases accept the protons readily, they are called strong bases. Strong acids completely react to their conjugate bases, while weak acids do not react to completion in water and form equilibrium [1] .

When strong acids are put into water, they go immediately to H_3O^+.

Therefore, an acid like HCl exists in water as H_3O^+.

A way of measuring how acidic or how basic a substance is the pH scale (short for " power of hydrogen")[1]. The pH scale is a logarithmic scale that measures the concentration of H_3O^+ in a solution according to the formula pH = -log [H_3O^+][1]. So, the higher the concentration of H_3O^+, the lower the pH value, and the more acidic the solution is. At 25° Celsius, a pH of 7 is neutral, pH values from 1to 7 are acidic, and pH values from 7 to 14 are basic[1]. A way to determine pH experimentally is to use a universal indicator, which is a substance that will change colour at a certain pH. If the right indicator is picked, equivalence point and end point should be the same. On the other hand, if the indicator was chosen poorly, it will change colour at some point substantially different from the equivalence point.

Due to the fact that most acid-base solutions are colourless it is very difficult to determine when a reactant has been totally consumed. That being said, indicators are utilized to make the equivalence point apparent. Indicators are weak acids or bases that are usually organic. The point at which an indicator is half way through its colour change is called the end point or transition point[4]. The concentration can thus be determined from the volume of base consumed at this point. When titrating a strong acid with a strong base any indicators can be used. Whereas, when titrating a weak acid with a strong base an indicator such as Phenolphthalein is used. On the other hand, when titrating a strong acid versus a weak base Bromothymol Blue can be used.

The reason these indicators are used is due to the pH at equivalence point. Strong base and a weak acid have a pH greater than 7. In contrary a strong acid with a weak base has a pH less than 7 at equivalence point [5].

Equations:

$$HA + OH^- \rightleftharpoons A^- + H_2O \qquad\qquad (1)$$

$$HIn + OH^- \rightleftharpoons H_2O + In^- \qquad\qquad (2)$$

According to the Bronsted Lowry theory. This can be explained with equation 1 and 2. Equation 1 is the reaction that happens when a strong acid and strong base react together. Once all of the hydrogen ions dissociate from the chlorine ions in hydrochloric acid, the base then pulls hydrogen ions from the indicator (shown in Equation 2). Once the indicator has the hydrogen ion taken from it, it changes colour, signifying the endpoint of the reaction between hydrochloric acid and sodium hydroxide. As expressed by the Ostwald theory colour changes our expected. The theory states that colour change is a result of ionization of an indicator.

A commonly employed procedure used to ascertain the concentration of an unknown acid or an unknown base in solution is referred to as titration. Titration is a vital skill to understand. The reason behind this is because it takes place on a daily basis in all different types of environments. Within all aspects of society, titrations are utilized. The main goal of a titration to find the unknown concentration of one of the reactants. Acid base titrations is the slow drop by drop reaction of an acid with a base. The standard base solution is added from a burette to an accurately measured volume, or aliquot, of

acid solution until the equivalence point, or stoichiometric point, is reached. Within a titration, the equivalence point is where the amount of titrant added completely neutralizes the solution. In other words, the equivalence point occurs when the mole ratio in the reaction exactly equals the mole ratio required by the stoichiometry of the reaction. At this point, neither the acid nor the base is in excess or is the limiting reagent. Additionally, the pH of the solution depends on the relative strength of the acid and strength of the base used in the titration. A monoprotic acid (HCl) is an acid that donates one proton and therefore has one equivalence point. Whereas polyprotic acids have an equivalence point per removal of proton. It is crucial to be able to see the equivalence point within a reaction. For this, an acid-base indicator, which changes colour near the equivalence point, is used. Sometimes pH can be measured with an instrumentally (instead of chemically with an indicator). A pH meter is used for this and measures a change in voltage which is caused by the change in H+. At the endpoint, a huge jump in pH/voltage occurs.

Hydrolysis is another important concept in this lab. Understanding the salt is very important due to the importance it plays within our body. Hydrolysis is the term used to describe the dissociation (splitting into ions) of a salt (an ionic compound) in water [6] . In most cases, ions will react with water in the solution. The ions that do not react with water are called " spectator ions". Spectator ions come from the first two groups of the periodic table, (alkali and alkaline eat metals) and the conjugate bases of strong acids. Spectator ions do not affect the pH of the solution. (1) All other ions react with water and either forms an acidic or basic solution. If the ions react with water and

produce H_3O^+ then the solution will become acidic and if it produces OH^-, the solution will become basic [1].

The purpose of this lab is to titrate a hydrochloric acid solution of unknown concentration and an acetic acid solution (vinegar) with standardized 0. 111M sodium hydroxide to determine the molarity of both solutions and the percentage composition of vinegar. As well, to observe the equivalence point between these two reactions.

Procedure:

The procedure that was listed in the First-Year Chemistry Lab Manual for Experiment S was precisely followed with a few minor adjustments that are discussed below.

The first portion of this lab is to determine the equivalence point of hydrochloric acid and acetic acid with sodium hydroxide 0. 111M. This process will be observed by using a pH meter to measure the ph. Firstly, within the experiment, deionized water was filtered through the 50 mL burette as well as 10mL of sodium hydroxide to clean the device and avoid contamination. Then, the burette was clamped to the retort stand and the basic solution sodium hydroxide was added to the zero mark. Air bubbles were removed from the burette to avoid inaccurate measurements. After that, the initial volume was recorded. Then, 10. 00mL of hydrochloric acid was pipetted into a beaker containing approximately 20mL of deionized water. Two drops of the appropriate indicator, bromothymol blue, was then added to the beaker that is then placed on the magnetic stirrer with 400

rpm. Then, the magnetic stirrer flea was placed in the beaker to merge the solutions together. The electrode was cautiously placed to avoid any damage to the electrode. The initial pH was recorded before adding sodium hydroxide slowly from the burette. At the beginning of the titration process, 1. 0mL of sodium hydroxide was continuously added to the beaker. After each interval the pH was recorded. When the pH started to fluctuate drastically, 0. 2mL was added instead of 1. 0mL each interval still recording the pH. This action was done to avoid missing the equivalence point. When the colour turns permanent blue, the endpoint has been obtained. In the event that the equivalence point is missed, continue the process but make sure multiple recordings are gathered after the endpoint to produce an accurate graph. The process was repeated using 5. 00mL acetic acid (CH $_3$ COOH) instead of using 10. 00mL of hydrochloric acid.

Next, the endpoint was measured rather than the equivalence point. The endpoint of the reaction was observed using the three indicators provided. This process was carried through using both sodium hydroxide and acetic acid. First, the process of sodium hydroxide. A 50. 00mL burette was gathered and rinsed thoroughly. Then 10. 00mL of sodium hydroxide was filtered through the burette for accuracy. The burette was then clamped to the retort stand and filled with sodium hydroxide to the zero mark. The initial volume was recorded. Then, 10. 00mL of hydrochloric acid was pipetted into three clean Erlenmeyer flasks. Next 2 drops of phenolphthalein were added to the first beaker and placed under the burette. Sodium hydroxide was slowly maneuvered into the flak until the endpoint was obtained. After the endpoint was reached, the final volume was recorded. Then, the process was

repeated using the other two indicators which were bromothymol blue and bromophenol blue. This whole procedure was carried through again utilizing 5. 00mL of acetic acid.

Finally, after observing are results, the correct indicator was selected, and titration was repeated for accuracy. The indicator chosen for hydrochloric acid was bromothymol blue and for acetic acid it was phenolphthalein.

Calculations:

See the attached Lab S Report Sheets for the Data and Calculations.

Discussion:

Acids and bases are substances capable of donating or accepting hydrogen ions. Hydrochloric acid, as illustrated in equation 3, is a strong acid that donates 100% of its protons. Due to the fact that hydrochloric acid is a strong acid, when put in water hydrochloric acid reacts immediately to H_3O^+. H_3O^+ is the strongest acid that exists within water and the difference cannot be obtained. This is called the levelling effect.

$$HCl + H_2O \rightleftharpoons Cl^- + H_3O^+ \quad (3)$$

Acetic acid is a weak acid as seen in equation 4 and has a conjugate acid-base pair related to one another through the gain or loss of one proton. A conjugate base as seen below has one less proton than its conjugate acid. Elsewise, a conjugate acid has one more proton than its conjugate base. Acetic acid dissociates partially into its conjugate base form. Due to the fact that acetic acid doesn't donate 100% of its protons it forms an equilibrium

reaction. In all acid base equilibria, the reaction always goes from strong to weak; thus, the weak side is always favoured. Acetic acid is weak therefore has a strong conjugate base.

$$CH_3COOH + H_2O \rightleftharpoons CH3COO^- + H_3O^+ \quad (4)$$

According to the data the purpose was achieved. Within the experiment, the molarity of HCl and CH_3COOH was determined at both equivalence and endpoint by using the filtration process. The molarity of hydrochloric acid at the endpoint was 0. 124M and at the equivalence point it was found to be 0. 121M. In contrary, the molarity of acetic acid in regard to the endpoint was 0. 291M and at the equivalence point was 0. 284M. The data collected is expected when taking the Bronsted Lowry theory into consideration.

As demonstrated by plotting pH versus volume of NaOH the graphs of both acetic acid and hydrochloric acid illustrate an s curve. (See graph 1 and 2) The reason it is an s-curve is because at first no NaOH substance has been added and is predominately made up of H_3O^+. When NaOH is added the composition of H_3O^+ slowly progresses into OH^- ions due to the NaOH dissociation. This process constantly occurs until the equivalence point is obtained, resulting in an inflection point on the graph. The inflection point is known as the equivalence point. At the equivalence point, the moles of NaOH added equals the moles of HCl. The composition of H_3O^+ ions are completely neutralized by OH^- ions. Therefore, the solution will be left with salt (NaCl) and water; thus, the leaving the solution with a neutral pH. After the equivalence point and while NaOH is still being added, the pH becomes

basic due to the fact that the HCl has completely neutralized. All that is now left are OH ⁻ ions and therefor the solution is basic. PH is only partially related to the strength of the acid. It is also dependent on the original concentration of the acid. In the study of finding the equivalence point when titrating HCl (a strong acid) with NaOH (a strong base) the pH at equivalence point is 7. 00. This result is anticipated due to the fact that the salt produced is a neutral hydrolysis salt. On the other hand, while titrating acetic acid with NaOH the pH is 8. 41. This result is expected because the pH at equivalence is greater than 7. The salt produced does a basic hydrolysis reaction.

Within this experiment different indicators were utilized, which resulted in multiple colour changes to occur. With the use of the indicator bromophenol blue, the colour initially was yellow and developed into a violet colour. Then, with bromothymol blue the colour changed drastically from yellow to light blue. Finally, when utilizing phenolphthalein, the colour gradually changed from a clear solution to light pink. Bromothymol blue has a pH range of 6. 0-7. 6 therefore, should be used when titrating a strong acid and a strong base. This indicator was chosen within the experiment when titrating hydrochloric acid and sodium hydroxide. Likewise, phenolphthalein has a pH range of 8. 4-10. 0 which is utilized when titrating a strong base with weak acid. Moreover, this indicator was used in the event of titrating acetic acid with sodium hydroxide. When first adding the sodium hydroxide the colour scarcely changed however, as more solution was added the colour drastically changed colour due to the equivalence point being reached. The colour changed is expected due to the Ostwald theory.

This experiment has many sources of error. Firstly, the molarities of the given solutions may not have been exact. This error would result in the calculations to be drastically different. Next, when titrating the hydrochloric acid with sodium hydroxide, the pH jumped from 2. 47 to 9. 92. Resulting in a very big gap in the plot created and uncertainty in determining the volume of sodium hydroxide added to reach the equivalence point. Additionally, parallax error occurred. Parallax error is the error that occurs when determining if the meniscus should be above, in the middle or at the bottom of the line on the glassware. The result of this is measurements are not as precise as desired and increase the uncertainty by a vast amount. Lastly, the nature of the reactant is a source of error. Substances can cling to the glassware therefore, when transferring from one glass to another, a small portion will remain in initial glass. One way to improve the error is to obtain more precise glassware.

Conclusion:

After completing the lab, it was discovered that the molarity of hydrochloric acid was 0. 124M using visual indicators and 0. 124M with a pH meter. As well the molarity of acetic acid was 0. 291M when using indicators and 0. 284M when using a pH meter. The pH at equivalence point of HCl was 7 and for CH_3COOH it was 8. 41. This lab was a success and has helped me develop a greater understanding towards acid-base titrations.

References

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