Abstract

Introduction

Purpose

The purpose of this experiment was to determine the extent to which the pH of a 150 mL buffer solution between Acetic Acid, $C_2H_4O_2$, and Sodium Acetate, $C_2H_3NaO_2$, varied with temperature to determine how it, transitively, affected the strength of the acid.

Background

The pH of a solution is a measure of the molar concentration of hydrogen ions in the solution; threby, being a measure of the acidity or basicity of the solution. A pH of less than 7 is basic, greater than 7 is acidic, and 7.0 is neutral. The pH of a solution can be mathematically represented as the negative logarathim of the Molar concentration of the hydrogen ions present in solution, as shown below.

$$pH = -\log_{10}[H^+]$$

The pH of a solution essentially states the extent of the veracity of an acid or base. The closer the pH of an acid is to 1, the stronger and more volatile the acid. Conversely, the closer the pH of a base is to 14, the stronger and more volatile the base.
The pH of a solution is also known to vary by temperature. In a study conducted by Ashton and Geary, it was found that the pH of a solution varied by temperature, depending on the initial measured pH at room temperature, as shown in the table below.

	Temperature		
pH Range	0°C	25°C	60°C
Acid	pH 0.99	pH 1.00	pH 1.01
Neutral	pH 7.47	pH 7.00	pH 6.51
Basic	pH 14.94	pH 14.00	pH 13.02

Figure 1: Relationship of pH and temperature (Ashton and Geary, 2005).

The pH of a solution varies with concentration as well. There is a linear relationship between the concentration of a solution and its subsequent pH. This is because the pH of a solution is the measure of the hydrogen ion concentration of the solution. For every increase in pH by a factor of 1, the concentration increases by a factor of ten.
The acid dissociation constant, K_a, is the measure of the strength of an acid in solution. The K_a is found by solving the expression for the following acid dissocation reaction:

$$\mathrm{HA}_{\,(aq)} \rightleftharpoons \mathrm{H}^+_{(aq)} + \mathrm{A}^-_{(aq)}$$

Where HA is the generic acid, H^+ is the hydrogen ion, and A^- is the conjugate base of the acid. The above reaction is in equilibrium when the concentrations of all the elements in the reaction is constant. The acid dissocation constant is therefore the products over the reactants as shown below:

$$K_a = \frac{[H^+][A^-]}{HA}$$

From the acid dissociation constant, the pK_aT of the acid can be derived. The pK_a of an acid states the acidity of a given hydrogen atom of excatly one molecule of that acid. The pK_a of an acid is essentially the pH at which it is exactly half dissociated. The pK_a of the acid can be mathematically calculated as the negative logarithm of the acid dissociation constant, as shown in the equation below.

$$pK_a = -\log_{10} K_a$$

The larger the value of the pK_a, the lesser the dissociation of the acid, thereby indicating a weak acid. The smaller the value of the K_a, the weaker the acid. Therefore, the smaller the value of the pK_a, the greater the dissociation of the acid, indicating a strong acid. The larger the value of the K_a the stronger the acid.

A buffer solution is a solution which consists of a weak Bronsted acid and its conjugate base, or a weak Bronsted base and its conjugate acid. Buffer solutions are quite resistant to pH changes, when small quantities of acid are added, as well.

There are two types of buffer solutions: Acidic and Alkaline. Acidic buffer solutions have a pH less than 7, and are composed of weak acids and its conjugate base. Alkaline buffer solutions have a pH greater than 7, and are composed of weak bases and its conjugate acid.

Buffer solutions essentially work by removing any hydrogen or hydroxide ions which might be added to it, thereby not changing the pH of the solution when an acid or base is added to it.

Buffer solutions are paramount to industry due to the inumerable practical applications which are present. One such industry is pharmaceuticals. in the Pharmaceutical industry, many therapeutic drugs are synthesizied to form buffer solutions to increase the shelf-life of the drugs, ensure the stability of treatments, maintaing the drug at a near neutral, constant pH to avoid irratition with skin, and much more. Buffer solutions are used in fermentation reactions, such as beer and yogurt, to ensure that there are no harsh changes and to achieve maximum yield. Buffers are also used in the manufacture of glue to ensure that there aren't changes in the highly sensitive chemical gelatine. Lastly, buffers are used in the soap industry to create soap with a pH of 5.5, the pH of our skin, and to avoid any irritation with our skin.

In a study conducted by Gerardo Gomez, Michael J. Pikal, and Nair Rodriguez-Hornedo, the pH changes of buffers were measured when Sodium Phospate buffers were induced by Salt percipitation during various far-from-equilibrium freezing temperatures. It was concluded that the greatest variations in pH occured at lower freezing temperatures around -10^oC. There was a linear correlation between colder temperatures and greater variations, as the temperature was closer to -10^oC, the greater the variation in pH from the buffer's initial equilibrium value.

This begged the question: How would the pH of a buffer soultion vary with an increase in temperature. This study was not only conducted to answer that very question, but also conducted to explore the practical implications involved with understanding the relationships of buffers and its subsequent pH. After determining the pH the effects on the K_a of the acid can be explored, thereby determining how an increase in the temperature of a buffer can affect the stregth of an acid. This knowledge can help various companies determine the proper temperature in which their buffers must be created, thereby maximizing profit, time, and the accuracy of the pH of the buffer.

Hypothesis

If an increase in the temperature of an Acid increases its pH; thereby increasing the decreasing the strength of the acid, then an increase temperature of the buffer solution between Acetic Acid, $C_2H_4O_2$, and Sodium Acetate, $C_2H_3NaO_2$, will cause an increase in temperature; thereby, weakening the acid.

Safety Information

Safety is paramount, and good laboratory practices should not only be informed, but also practiced. The risk of burn was apparent, all throughout the experiment. Simple measures must be taken to prevent any sort of burn injuries. The hot-plates used, and hot beakers should be stationed at an area, away from foot traffic. Proper equipment should be used to touch and move the hot equipment, and lastly, safety googles must be worn at all times, in case of splattering.

Materials

1. Acetic Acid (50 mL)
2. Sodium Acetate (15 g)
3. 50 mL Beaker (3)
4. 150 mL Beaker (1)
5. pH Probe (1)
6. Temperature Probe (1)
7. Hot Plate (1)

Methods

Procedure

Initially, 100 mL of Acetic Acid was measured with a pipette and poured into a 150 mL beaker. The 100 mL Acetic Acid was then diluted down to 0.11 M by pouring 10mL of water into 90mL of Acetic Acid. The pH of the 0.11 M Acetic Acid was then measured to 3.72. Next, 10.1 g of Sodium Acetate was weighed, and poured into the Acetic Acid to create the buffer solution. THe mixture was then stirred vigorously for approximately 3 minutes until all the Sodium Acetate was dissolved. The pH of the buffer solution was then measured to 6.29. Next, the buffer solution was placed on a hot plate, and a temperature probe was inserted to record the temperautre. The pH of the buffer solution was recorded during every increase of 10^oC until it reached 102.6^oC. The $K_a$and pK_a values were then calculated and analyzed.

Results

Raw Data

Initial pH of Acetic Acid and Buffer Solution

Substance	pH
Acetic Acid	3.72
Buffer Solution	6.29

Temperature vs pH of Buffer

Temperature (^oC)	pH
22.2	6.29
32.5	6.28
42.5	6.26
52.5	6.25
62.5	6.23
72.5	6.22
82.5	6.20
102.5	6.05

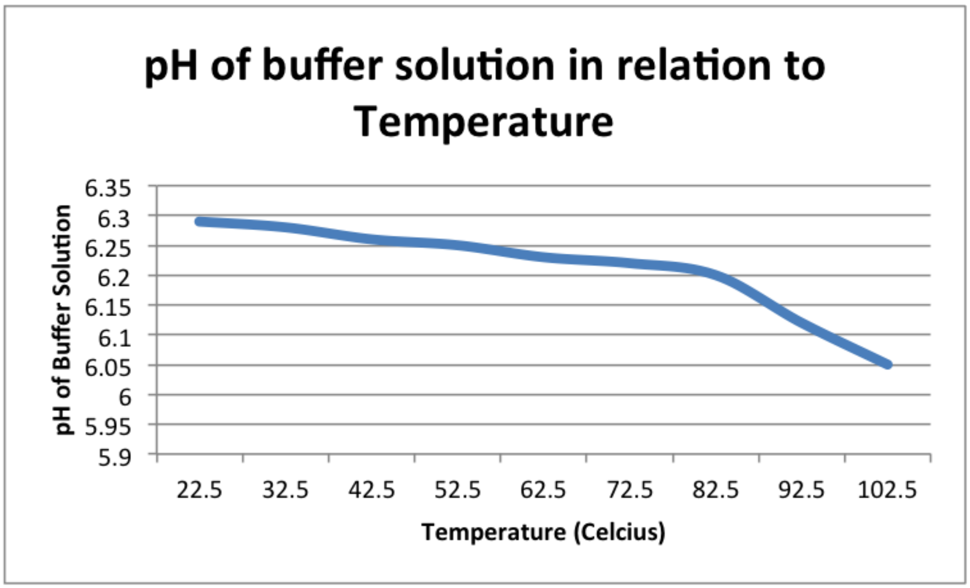

Figure 2: Graph of Temperature vs pH of Buffer

Calculations

Discussion

Conclusion

Experimental Error

Improvements