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.

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 (2005, Ashton and Geary).

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 as well.

Hypothesis

Safety Information

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

Results

Raw Data

Calculations

Discussion

Conclusion

Experimental Error

Improvements