

Exploring the Effect of Temperature on the Concentration of Salicylic Acid in Acetylsalicylic Acid

Rojina Karimirad

Bayview Secondary School

Introduction

Scope

Aspirin is used for a variety of reasons, from reducing fever to providing relief for mild pain associated with headaches, menstrual cramps, and arthritis (*Aspirin: MedlinePlus Drug Information*, n.d.). Its versatility leads to aspirin tablets being present in almost every home, which makes the consequences of their mishandling that much more impactful. Generally, patients are instructed not to store aspirin in areas exposed to excess heat and humidity as they cause aspirin to break down into salicylic acid and acetic acid (*Aspirin: MedlinePlus Drug Information*, n.d.). Salicylic acid is the substance responsible for aspirin's medicinal properties, and fluctuations in its concentration can lead to dosage inaccuracies. This leads to questions and concerns about the correlation between temperature and the concentration of salicylic acid within hydrolyzed aspirin.

Background Information

The active ingredient in aspirin tablets is acetylsalicylic acid, a salicylate. Salicylates are esters of salicylic acid, the substance that aspirin owes its anti-inflammatory properties to (*Aspirin: MedlinePlus Drug Information*, n.d.). Acetylsalicylic acid is therefore, along with acetic acid, one of the products of the esterification of salicylic acid with acetic anhydride. Once an aspirin tablet is ingested and enters the small intestine, the acetylsalicylic acid undergoes a hydrolysis reaction, where acetylsalicylic acid reacts with water to break the ester bond and produce salicylic acid and acetic acid, as shown in *Figure 1* (Awtry & Loscalzo, 2000, The Editors of Encyclopaedia Britannica, 1998).

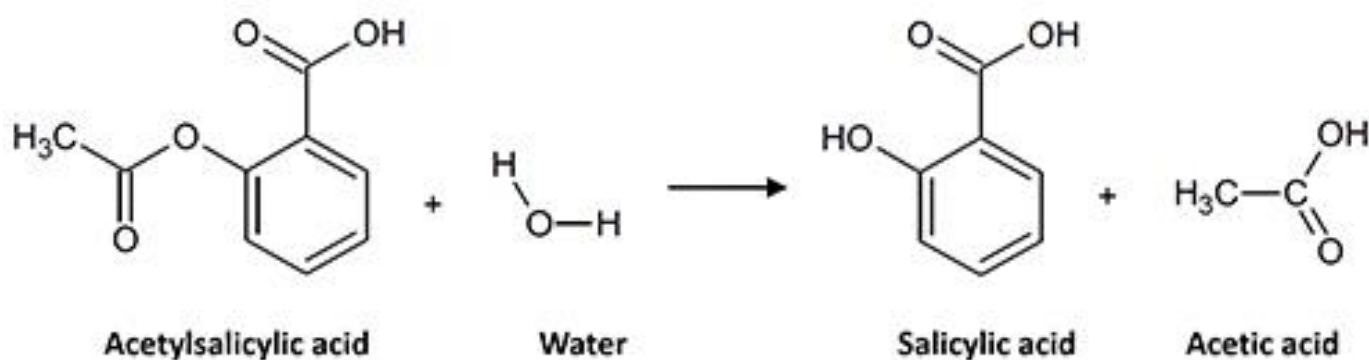


Figure 1: Diagram displaying the hydrolysis of acetylsalicylic acid into salicylic acid and acetic acid (Marino, 2021)

Since salicylic acid is the reason for aspirin's medicinal properties, its concentration in hydrolyzed acetylsalicylic acid is incredibly important, and the effect of varying temperature on this concentration must be explored to provide patients with adequate handling instructions. According to Collision Theory, rate of reaction, in this case hydrolysis, increases with temperature as higher temperatures increase the rate of collision and the proportion of successful collisions in a reaction (Key, 2014). This leads to the hypothesis that the concentration of produced salicylic acid increases as temperature increases. Since temperature and rate of reaction have an exponential relationship, it would be expected for the increase in concentration of salicylic acid to occur exponentially as well (*Arrhenius Kinetics Analysis*, n.d.).

Once salicylic acid, $C_7H_6O_3$, is dissolved in water, it produces, with a 1:1 mole ratio, the salicylate dianion, $[C_7H_4O_3]^{2-}$, which is the conjugate base of salicylic acid. When iron (III) nitrate, $Fe(NO_3)_3$, dissolves in water, it disassociates into $[Fe]^{3+}$ and $[NO_3]^-$ ions. The iron (III) ions then react with the water to produce the hexaaquairon (III) ion, $Fe(H_2O)_6^{3+}$, again with a 1:1 mole ratio. The salicylate dianion and hexaaquairon (III) ions finally react in the solution, with a 1:1 mole ratio, to form tetraaquosalicylatriron, a violet-coloured complex that absorbs green light.

The concentration of salicylic acid in a solution can therefore be measured based on the intensity of the colour of the violet complex, using a spectrophotometer and the Beer-Lambert Law. The Beer-Lambert Law states that the absorbance of a substance is directly proportional with its concentration, with the equation, $A = \epsilon lc$, where A is absorbance, ϵ is the molar extinction coefficient, l is the length of the light path, and c is concentration in mol/dm^3 ("The Beer-Lambert Law", 2023). The molar extinction coefficient is "a measure of how strongly a chemical species or substance absorbs light at a particular wavelength", and is different for every substance (*What Is a Molar Extinction Coefficient?* / AAT Bioquest, n.d.). The value of the molar extinction coefficient of salicylic acid can be found via a calibration curve drawn using known concentrations of salicylic acid, where the x axis represents concentration, and the y axis represents absorbance (Prichard, 2003). According to the equation of the line, $y = mx + b$, the coefficient of x , m , is the slope. So, in a graph of absorbance vs. concentration, the coefficients of c in the Beer-Lambert equation, ϵl , would be the slope. Therefore, to calculate for ϵ , the slope of the absorbance vs. concentration graph would need to be divided by l . The ϵ value could then be used in the actual experiment to solve for the unknown concentrations of salicylic acid produced at different temperatures.

As previously mentioned, the above process requires calibration using a set of known concentrations of diluted salicylic acid in iron (III) nitrate solution. The two substances react with a 1:1 mole ratio, but since the goal of the calibration is for the salicylic acid specifically to fully react, it would be safer to have an excess of iron (III) nitrate. This is shown in the calculations below, which are based on the masses and volumes used in Solution 5 of the calibration process. Calculations are being shown for Solution 5 because it contains the highest concentration of salicylic acid, so if iron (III) nitrate is in excess in Solution 5, it is in excess for the four other solutions as well.

$$n_{\text{salicylic acid}} = \frac{0.07g \pm 0.01}{138.13g/mol^*} \times (0.0250dm^3 \pm 0.0001)$$

$$n_{\text{salicylic acid}} = \frac{0.00051mol \pm 0.00007}{0.0200dm^3 \pm 0.0001} \times (0.0250dm^3 \pm 0.0001)$$

$$n_{\text{salicylic acid}} = \frac{0.00051mol \pm 10\%}{0.0200dm^3 \pm 0.5\%} \times (0.025dm^3 \pm 0.4\%)$$

$$n_{\text{salicylic acid}} = 0.000064mol \pm 10\%$$

$$n_{\text{salicylic acid}} = 0.000064mol \pm 0.000006$$

$$n_{\text{iron (III) nitrate}} = 0.0125mol/dm^3 \times (0.0475dm^3 \pm 0.0001)$$

$$n_{\text{iron (III) nitrate}} = 0.000594mol \pm 0.000001$$

$$\therefore 0.000594mol \pm 0.000001 > 0.000064mol \pm 0.000006 \therefore \text{iron (III) nitrate is in excess}$$

*Molar mass calculated based on values obtained from the "Chemistry Data Booklet", 2001.

The different temperature values for the independent variable were chosen based on the scope of this exploration, which was to explore temperature's effect on the production of salicylic in the context of household usage. Selecting extremely high temperature values would be unrealistic for a household, which is why the lowest temperature value was chosen to be 30.0°C, resembling the temperature of a hot summer day. 30.0°C was also chosen because it was higher than the average room temperature of 20°C and the exact temperature could therefore be achieved through the use of a hot plate (*What Is the Ideal Room Temperature?*, n.d.). For reference, the room temperature of the laboratory on the first day of experimentation was recorded to be 19.8°C. The independent values increase by larger increments of 15.0°C, so they significantly affect the dependent concentration values and lead to a clearer correlation. The highest independent value is 90.0°C, which is the maximum allowable wood cabinet temperature, again being realistic for where aspirin would be stored in a household (*NE63T8111SS*, n.d.).

The pH of the acetylsalicylic acid solution has an effect on the rate of hydrolysis of aspirin, with pH2 leading to the highest rate of reaction (Al-Khalaf et al., 2019). Since only the correlation between temperature and hydrolysis is being investigated in this exploration, the pH level of the acetylsalicylic acid solution is controlled using pH7 buffer solution. Buffer solutions contain weak conjugate acid-base pairs and are therefore able to neutralize small amounts of acidic and basic additions to the solution, resisting a change in pH ("Introduction to Buffers", 2023).

Research Question

What is the effect of temperature (30.0°C, 45.0°C, 60.0°C, 75.0°C, 90.0°C) on the concentration of salicylic acid in a fixed mass of acetylsalicylic acid at a controlled pH, as measured after 3:00 minutes by the absorbance of the solution at 521nm using a spectrophotometer?

Variables

Independent Variable: Temperature of the acetylsalicylic acid solution (30.0°C, 45.0°C, 60.0°C, 75.0°C, 90.0°C), manipulated using a hot plate and measured using a digital thermometer

Dependent Variable: Concentration of salicylic acid in the solution, calculated based on the Beer-Lambert equation and the absorbance of the solution at 521nm as measured by a spectrophotometer

Controlled Variables:

- Concentration of acetylsalicylic acid in each trial was $0.006\text{mol/dm}^3 \pm 0.001$, as 0.05g of acetylsalicylic powder was mixed into 50cm^3 of ethanol and pH7 buffer solution
- Concentration of iron (III) nitrate was 0.0125mol/dm^3 throughout all trials and the calibration process
- Concentration of salicylic acid was $0.026\text{mol/dm}^3 \pm 0.003$ throughout the calibration process
- Volume of acetylsalicylic acid solution remained 50.0cm^3 throughout all trials, with 12.5cm^3 of ethanol and 37.5cm^3 of pH7 buffer solution
- Volume of reacting solution in each trial was 10.0cm^3 throughout all trials, with 9.0cm^3 of iron (III) nitrate solution and 1.0cm^3 of acetylsalicylic solution
- Duration of hydrolysis of acetylsalicylic acid remained 3:00 minutes after the powdered acid had disappeared throughout all trials
- pH of acetylsalicylic acid solution was controlled using pH7 buffer solution

Methodology

Materials

- Salicylic acid powder ($0.35 \pm 0.01\text{g}$)
- Acetylsalicylic acid powder ($1.00 \pm 0.01\text{g}$)
- 0.0125mol/dm^3 iron (III) nitrate solution ($427.5\text{cm}^3 \pm 0.1\text{mL}$ or $\pm 0.05\text{mL}$)
- Ethanol ($255.0\text{cm}^3 \pm 0.1\text{mL}$ or $\pm 0.05\text{mL}$)
- pH7 buffer solution ($765.0\text{cm}^3 \pm 0.1\text{mL}$)

Apparatus

- Hot plate ($\pm 1^{\circ}\text{C}$), 1
- Electronic balance ($\pm 0.01\text{g}$), 1
- Spectrophotometer (± 0.001 for absorbance), 1
- Digital thermometer ($\pm 0.1^{\circ}\text{C}$), 1
- Ruler ($\pm 0.05\text{cm}$), 1
- Timer ($\pm 0:01$), 1
- Glass cuvette, 1
- 50mL beaker, 26
- 100mL beaker, 20
- 25mL graduated cylinder ($\pm 0.1\text{mL}$), 26
- 10mL graduated cylinder ($\pm 0.05\text{mL}$), 26
- Stirring rod, 21
- Scoopula, 21
- Pipette, 25
- Funnel, 25
- Weighing boat, 21

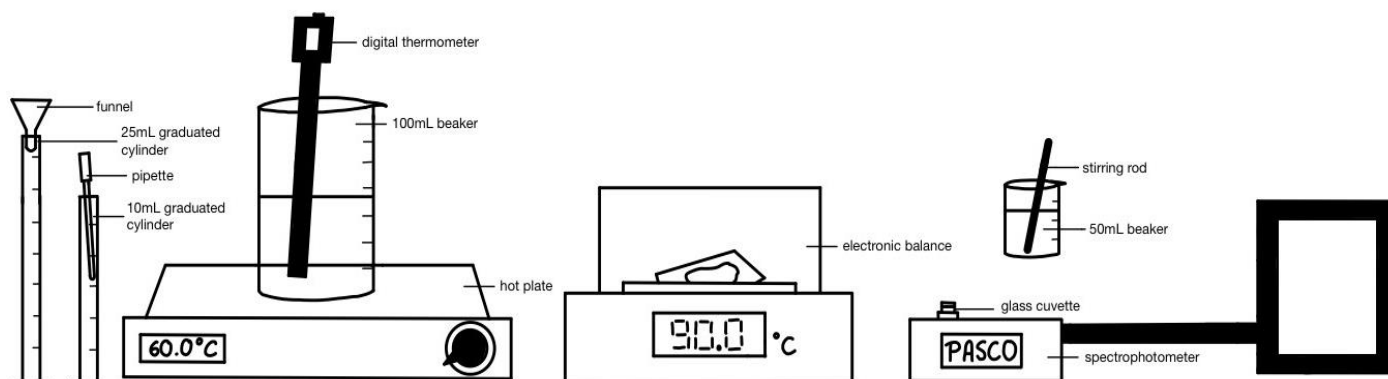


Figure 2: Depicted process of making and measuring the absorbance of each solution

Procedure

Preparation of Salicylic Acid Stock Solution

1. 5.00cm^3 of ethanol was added to 15.0cm^3 of pH7 buffer solution in a 50mL beaker.
2. 0.07g of salicylic acid was mixed into the solution to make salicylic acid solution.
3. The solution was put on a hot plate at around 240.0°C and stirred until the powder disappeared.

Calibration Curve

1. Spectrophotometer was blanked using distilled water, as water should not absorb any light.
2. 0.50cm^3 salicylic acid stock solution was added to 49.5cm^3 iron (III) nitrate solution in a 50mL beaker.
3. The solution was put into a glass cuvette using a pipette, ensuring all the bubbles were burst by tapping the cuvette.
4. The cuvette was placed into the spectrophotometer, fingers were kept away from its clear sides.
5. Absorption of the solution at 521nm^{**} wavelength was recorded via the spectrophotometer.
6. Steps 1-5 were repeated for 1.00cm^3 , 1.50cm^3 , 2.00cm^3 , and 2.50cm^3 of salicylic acid solution and 49.0cm^3 , 48.5cm^3 , 48.0cm^3 , and 47.5cm^3 of iron (III) nitrate solution respectively, and a calibration curve was drawn.

Experiment

1. In a 100mL beaker, 12.5cm^3 of ethanol was added to 37.5cm^3 of pH7 buffer solution.
2. The beaker was placed on a hot plate until it reached 30.0°C .
3. 0.05g of powdered aspirin was measured and added to the solution. The solution was stirred with a stirring rod until the powder had disappeared.
4. After the powder disappeared, the solution was left for 3:00 minutes for the hydrolysis to occur.
5. 1.00cm^3 of the solution was added to 9.00cm^3 of iron (III) nitrate solution in a 50mL beaker.
6. The new solution was put into a glass cuvette using a pipette.
7. The cuvette was placed into the spectrophotometer and the absorbance at 521nm^{**} was recorded.
8. Steps 1-7 were repeated for the temperatures 45.0°C , 60.0°C , 75.0°C , and 90.0°C .

***521nm was found to be the wavelength showing Solution 1's peak absorbance during a test run, hence why it was chosen. The green light that violet absorbs is considered to have a wavelength range of 495-570nm, which supports this choice (Helmenstine, 2020).*

Safety, Ethical, and Environmental Concerns

Safety: Goggles and gloves were worn throughout the experiment to prevent injuries involving glassware and avoid making contact with acetylsalicylic acid, salicylic acid, iron (III) nitrate, and ethanol as they can all irritate the eyes and the skin (CDC - NIOSH Pocket Guide to Chemical Hazards - Acetylsalicylic Acid, n.d., CDC - NIOSH Pocket Guide to Chemical Hazards - Iron Salts (Soluble, as Fe), n.d., Salicylic Acid SDS (Safety Data Sheet) / Flinn Scientific, n.d., ThermoFisher Scientific, 2022). Handling of ethanol was done inside a fume hood as it is a highly volatile substance and must not be inhaled (Ethanol (Ethyl Alcohol) - DCCEEW, n.d.).

Environmental: Solutions used were dilute and could therefore be disposed of down the sink with large amounts of water.

Ethical: There were no ethical concerns as no living organisms were used.

Data Collection

Table 1: Absorbance of iron (III) nitrate and salicylic acid solutions with increasing volume of salicylic acid and decreasing volume of iron (III) nitrate

Solution Number	Volume of Salicylic Acid Solution ($0.05 \pm \text{cm}^3$)	Volume of Iron (III) Nitrate Solution ($\pm 0.1 \text{cm}^3$)	Absorbance (± 0.001)
1	0.50	49.5	0.405
2	1.00	49.0	0.749
3	1.50	48.5	1.138
4	2.00	48.0	1.249
5	2.50	47.5	1.347

Table 2: Absorbance of iron (III) nitrate and acetylsalicylic acid solutions with increasing temperatures

Temperature ($\pm 0.1^\circ\text{C}$)	Absorbance (± 0.001)			
	Trial 1	Trial 2	Trial 3	Trial 4
30.0	0.542	0.538	0.544	0.539
45.0	0.578	0.576	0.580	0.575
60.0	0.597	0.604	0.598	0.603
75.0	0.614	0.623	0.618	0.619
90.0	0.635	0.642	0.639	0.641

Table 3: Colour and presence of precipitate observed in solutions of acetylsalicylic acid

Temperature ($\pm 0.1^\circ\text{C}$)	Qualitative Observations
30.0	- Yellow-orange, the colour of iron (III) nitrate solution - Precipitate formed after stirring
45.0	- Yellow-orange, the colour of iron (III) nitrate solution
60.0	- Yellow-orange, the colour of iron (III) nitrate solution
75.0	- Pale violet colour
90.0	- Light violet, but darker than the 75.0°C solution

Data Processing

Sample calculation for finding the concentration of salicylic acid in each solution in the calibration process. Showing the calculation for Solution 1 in Table 1.

$$\begin{aligned}
 c_{\text{salicylic acid}} &= \frac{\left(\frac{m_{\text{salicylic acid}}}{M_{\text{salicylic acid}}} \right) \times (V_{\text{stock solution used}})}{V_{\text{solution}}} \\
 c_{\text{salicylic acid}} &= \frac{\left(\frac{0.07 \text{g} \pm 0.01}{138.13 \text{g/mol}} \right) \times (0.00050 \text{dm}^3 \pm 0.00005)}{0.0200 \text{dm}^3 \pm 0.0001} \\
 c_{\text{salicylic acid}} &= \frac{0.00051 \text{mol} \pm 0.00007}{0.0200 \text{dm}^3 \pm 0.0001} \times (0.00050 \text{dm}^3 \pm 0.00005) \\
 c_{\text{salicylic acid}} &= \frac{0.00051 \text{mol} \pm 10\%}{0.0200 \text{dm}^3 \pm 0.5\%} \times (0.00050 \text{dm}^3 \pm 10\%) \\
 c_{\text{salicylic acid}} &= 0.00026 \text{mol/mol} \pm 0.00005
 \end{aligned}$$

The concentration of salicylic acid for all the solutions in *Table 1* can be seen in the table below.

Table 4: Absorbance of iron (III) nitrate and salicylic acid solutions with increasing concentrations of salicylic acid

Solution Number	Concentration of Salicylic Acid (mol/dm ³)	Absorbance (± 0.001)
1	0.00026 \pm 0.00005	0.405
2	0.0005 \pm 0.0001	0.749
3	0.00076 \pm 0.00008	1.138
4	0.0010 \pm 0.0001	1.249
5	0.0013 \pm 0.0001	1.347

The following graph was plotted based on the absorption values and the concentrations of salicylic acid calculated. The A-intercept of the linear regression was set to zero as a solution with no salicylic acid would have an absorbance of 0.

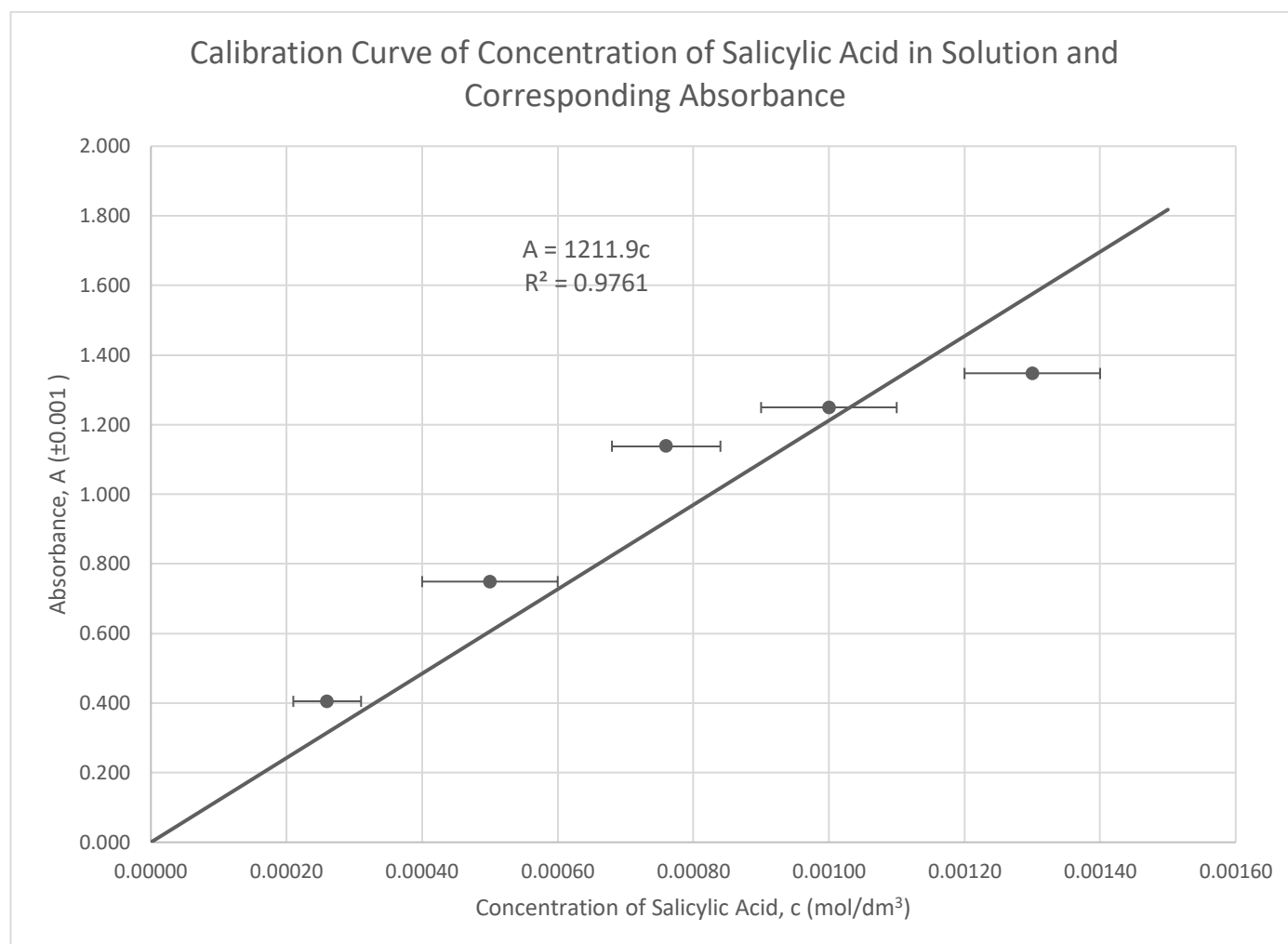


Figure 3: Calibration curve plotted using known concentrations of salicylic acid in solution and the absorption values recorded in Table 4, vertical error bars too small to be visible

The calculation for determining the molar extinction coefficient, ϵ , based on the equation of the linear regression in *Figure 3*, which shows that the slope of the line is 1211.9, can be seen below. The length of the light path of the cuvettes was considered to be from one inside edge of each cuvette to the other, which was measured to be $1.00\text{cm} \pm 0.05$ with a ruler.

Let m be the slope of the linear regression.

$$\epsilon = \frac{m}{l}$$

$$\epsilon = \frac{1211.9(\text{mol}/\text{dm}^3)^{-1}}{1.00\text{cm} \pm 0.05}$$

$$\epsilon = 1212(\text{mol}/\text{dm}^3)^{-1}\text{cm}^{-1} \pm 60$$

Based on the calculated value of the molar extinction coefficient and the Beer-Lambert equation, the value for the unknown concentrations of the solutions with varying temperatures can be calculated for each trial and then averaged out. The following sample calculation is for $30.0 \pm 0.1^\circ\text{C}$, Trial 1, with values obtained from *Table 4*.

$$c = \frac{A}{l\epsilon}$$

$$c = \frac{0.542 \pm 0.001}{(1.00\text{cm} \pm 0.05)(1212(\text{mol}/\text{dm}^3)^{-1}\text{cm}^{-1} \pm 60)}$$

$$c = \frac{0.542 \pm 0.2\%}{(1.00\text{cm} \pm 5\%)(1212(\text{mol}/\text{dm}^3)^{-1}\text{cm}^{-1} \pm 5\%)}$$

$$c = 0.00045\text{mol}/\text{dm}^3 \pm 0.00004$$

The calculated concentrations of salicylic acid for all solutions in *Table 4*, along with the mean values corresponding to each temperature, can be seen in the table below.

Table 5: Concentration of salicylic acid in acetylsalicylic acid with increasing temperatures of solution

Temperature ($\pm 0.1^\circ\text{C}$)	Concentration of Salicylic Acid (mol/dm^3)				
	Trial 1	Trial 2	Trial 3	Trial 4	Mean
30.0	0.00045 ± 0.00004	0.00044 ± 0.00004	0.00045 ± 0.00004	0.00044 ± 0.00004	0.00044 ± 0.00004
45.0	0.00048 ± 0.00005	0.00048 ± 0.00005	0.00048 ± 0.00005	0.00047 ± 0.00005	0.00048 ± 0.00005
60.0	0.00049 ± 0.00005	0.00050 ± 0.00005	0.00049 ± 0.00005	0.00050 ± 0.00005	0.00050 ± 0.00005
75.0	0.00051 ± 0.00005	0.00051 ± 0.00005	0.00051 ± 0.00005	0.00051 ± 0.00005	0.00051 ± 0.00005
90.0	0.00052 ± 0.00005	0.00053 ± 0.00005	0.00053 ± 0.00005	0.00053 ± 0.00005	0.00053 ± 0.00005

The following graph was plotted based on the temperature and mean concentration values in *Table 5*.

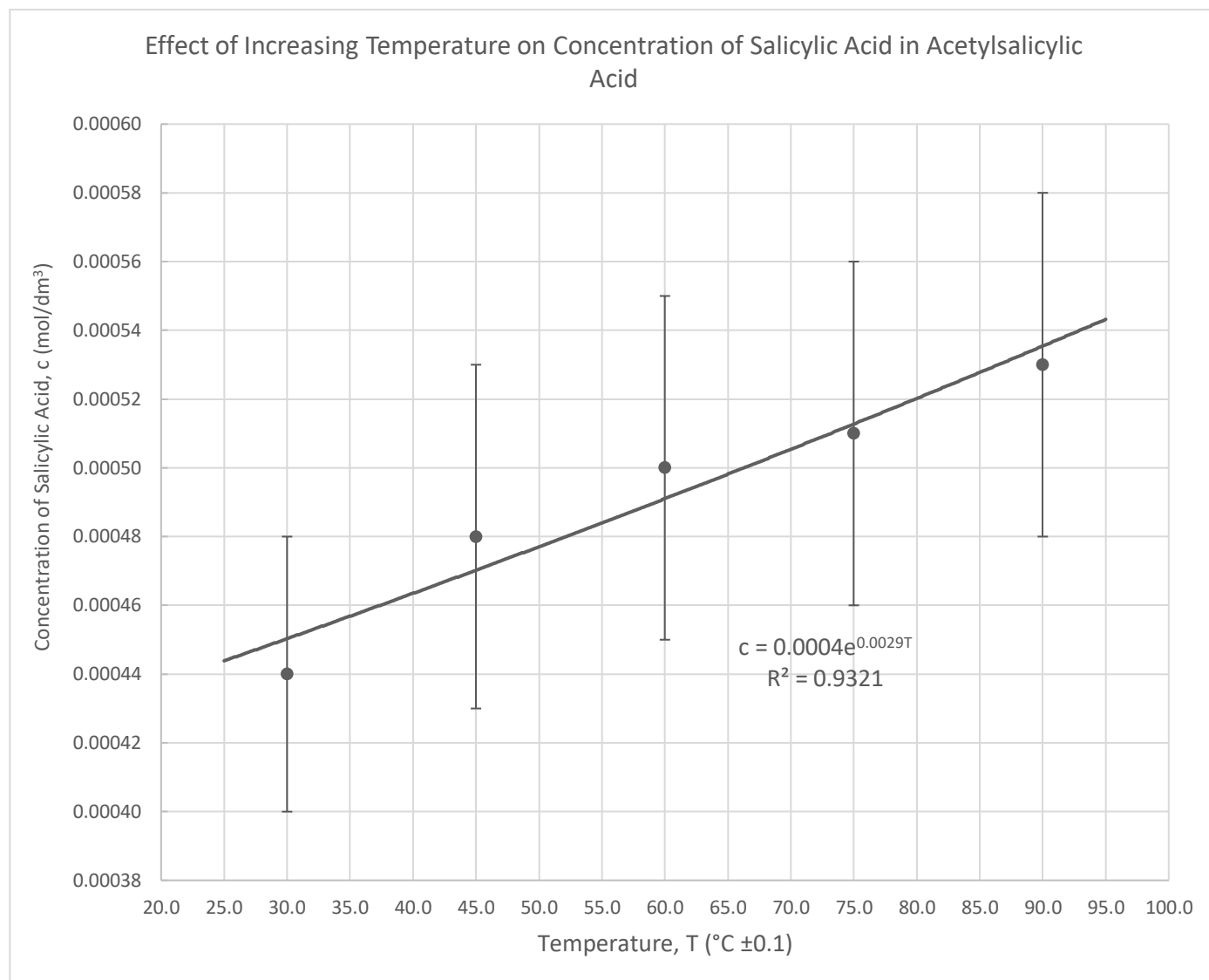


Figure 4: Graph displaying the correlation between temperature and concentration of salicylic acid in hydrolyzed acetylsalicylic acid based on values recorded in *Table 5*, horizontal error bars too small to be visible

Evaluation

Conclusion

The purpose of this experiment was to deduce the effect of increasing temperature on the concentration of salicylic acid in hydrolyzed acetylsalicylic acid. Based on the data calculated, shown in *Table 5*, and graphed in *Figure 4*, there is a directly proportional relationship between temperature and the concentration of salicylic acid in acetylsalicylic, as the concentration steadily increases as the temperature increases. This confirms the hypothesis that there is a positive correlation between temperature of acetylsalicylic acid and the concentration of salicylic acid produced in it through hydrolysis.

The graph plotted to demonstrate the relationship between temperature and concentration of salicylic acid, *Figure 4*, has a high R^2 value of 0.9321, showing that the positive correlation between the independent and dependent variables is in fact strong. The calibration curve, *Figure 3*, also has a high R^2 value, 0.9761, proving it reliable for being the basis of all other carried out calculations, though the data points seem to plateau at the two highest concentrations of salicylic acid as opposed to increasing in a linear fashion. The data from the experiment itself is very precise as the concentration values calculated for different trials of the same temperature do not vary by more than $0.00001\text{mol/dm}^3 \pm 0.00005$, as can be seen in *Table 5*. This strong positive correlation obtained is supported by Collision Theory, since an increase in temperature increases the average kinetic energy and proportion of successful collisions amongst particles and increases the rate of hydrolysis, which in turn results in the increasing concentration of salicylic acid recorded. The graph in *Figure 4*, however, does not align with the theory that temperature and rate of hydrolysis have an exponential relationship, as the regression model is nearly linear. This may be a limitation of having chosen relatively lower temperatures for the independent values due to the scope of the experiment, as the graph of $c = 0.0004e^{0.0029T}$ becomes very steep at around $c = 4.50000 \times 10^3 \text{mol/dm}^3 \pm 0.00005$, which is a very value compared to the calculated concentrations of salicylic acid. It may also be a result of systematic errors.

Sources of Systematic Error

Performing four trials of each temperature accounted for most random errors as seen in the high precision in the data, therefore the errors significantly affecting the results of the experiment were all systematic.

At the second highest concentration of salicylic acid in *Figure 3*, the graph reaches a plateau. This disruption in the linear increase in absorption suggests that at that point, not all of the salicylic acid reacted with the iron (III) nitrate, or in other words, there was not enough iron (III) nitrate for all of the salicylic acid to react. The iron (III) nitrate solution used throughout the experiment was made by dissolving iron (III) nitrate powder in distilled water. It initially dissolved in the water easily, being an ionic compound, but after being left in the lab for a few days, there was precipitate observed at the bottom of the bottle holding the solution. It is likely that this occurred due to slow hydrolysis of the iron (III) in the solution, which produced iron (III) hydroxide, an orange precipitate (Stefánsson, 2007). This would have reduced the concentration of iron (III) nitrate available in the solution for the salicylic acid to react with, explaining why despite iron (III) nitrate being in excess stoichiometrically, the salicylic acid did not fully react. This error can be prevented by the remaking of iron (III) nitrate solution in the beginning of every lab day, avoiding the slow formation of precipitate.

As shown in the qualitative observations written in *Table 3*, the powdered acetylsalicylic acid did not fully dissolve in solutions at 30.0°C . This is due to the smaller amount of energy available in lower temperatures, which makes it harder for the solvent, in this case ethanol, to break down the solute, in this case acetylsalicylic acid powder, and dissolve it. This error resulted in a larger difference between the concentration values of the lowest and second lowest

temperatures, 30.0°C and 45.0°C, as seen in *Table 5*, being one of the reasons why the *Figure 4* graph did not resemble one of exponential growth, since a characteristic of exponential growth is the differences in dependent value increasing as the independent value increases. A way of minimizing this error would be to select a higher starting temperature for the independent variable to provide the solution with enough energy to fully dissolve the acetylsalicylic acid powder.

Another reason for the resulting relationship between temperature and concentration being more linear is the solvent of acetylsalicylic acid, ethanol. Ethanol is a highly volatile substance, which results in it evaporating at a fast rate, especially as the temperature increases and provides it with more energy to change state. At higher temperatures, this fast evaporation of ethanol could have resulted in some of the acetylsalicylic acid being left undissolved, therefore not undergoing hydrolysis, and not producing the maximum yield of salicylic acid to react with iron (III) nitrate. This would again result in the concentration not demonstrating exponential growth as the temperature increases. If this issue were to be prevented by lowering the temperatures of the solutions, the previous source of error would be heightened. An alternate improvement would therefore be to substitute the solvent of acetylsalicylic acid with another substance entirely, a substance such as dimethylformamide or DMF. Similar to ethanol, DMF is capable of dissolving acetylsalicylic acid, but it has low volatility and would therefore not quickly evaporate at higher temperatures, fully dissolving the acetylsalicylic acid and allowing it to hydrolyze (Cayman Chemical Company, 2022, *Dimethylformamide (DMF)* / Eastman, n.d.).

Extensions to Experiment

Some aspirin pills have enteric coatings to keep the acetylsalicylic inside the pills intact and prevent the acidic conditions of the stomach from breaking down the pill before it reaches the small intestine (*Enteric Coated Aspirin Oral: Uses, Side Effects, Interactions, Pictures, Warnings & Dosing* - WebMD, n.d.). As an extension, the effect of different coatings on the hydrolysis of acetylsalicylic acid, the concentration of produced salicylic acid in solution, and the resistance of the pill to temperature could be explored.

Another possible extension would be to explore the kinetics of the hydrolysis of acetylsalicylic acid. The rate of the hydrolysis could be calculated by dividing the change in concentration of acetylsalicylic acid by the time values at the initial and final concentrations recorded. Since a buffer solution was used in the experiment, the pH level remained constant, resulting in the hydrolysis of aspirin being a first-order reaction (Onah, 2004). The rate of hydrolysis can therefore be written as $Rate = k[A]$, where k is the rate constant and $[A]$ is the concentration of acetylsalicylic acid. The rate constant can be obtained by plotting a graph of the natural logarithm of the concentration of acetylsalicylic acid, $\ln [A]$, vs. time, t , with the slope of the line being $-k$ (*Kinetics*, n.d.). The correlation between the rate constant and temperature of the solution can then be examined.

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