

LAB REPORT

2_ACDA6020_R_Dingwall

by Rebecca Dingwall

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Figure 1 Conical flask [13]

LAB REPORT 2

Rate of Reaction

ABSTRACT 1

The aim of this activity was to determine the rate law for the formation of iodine in the reaction between hydrogen peroxide and iodide ions in an aqueous solution, using vitamin C as a reducing agent and starch as an indicator.

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ACDA6020

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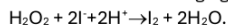
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1. Introduction

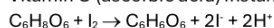
This lab report was required to gain a better practical understanding of kinetics and REDOX reactions employing the 'clock' method. The clock method, known by its dramatic change from colourless -to-blue is often employed in chemistry to explore reaction rate [1]. Traditionally the rate of reaction is found by measuring how the concentration of the reactant changes with time. With this experiment the gradual formation of iodine cannot be easily measured due to the immediate reduction of vitamin C until it is fully consumed.

1.1 Aim

The aim of this activity was to determine the rate law for the formation of iodine in the reaction between hydrogen peroxide and iodide ions in an aqueous solution, using vitamin C as a reducing agent and starch as an indicator *what does this indicator indicate?*



Vitamin C (ascorbic acid) instantly reduces iodine back to iodide.



1.2 Hypothesis

Based on previous experience and knowledge the assumption is that the rate law for both reactions is going to be either first order or second order. Not likely to be zero order due to the absence of a catalyst [2].

1.3 Variables

Trial 1: Order of reaction in H_2O_2

The dependant variable is the time taken for a colour change.

The independent variable is concentration of hydrogen peroxide.

The controls are potassium iodide, vitamin C (ascorbic acid), starch and water.

Trial 2: Order of reaction in I^-

The dependant variable is the time taken for a colour change.

The independent variable is concentration of potassium iodide

The controls are hydrogen peroxide, vitamin C (ascorbic acid), starch and water.

2. Theory

As previously mentioned in the introduction the basic principles of the iodine clock reaction are to give a better understanding of kinetics and REDOX reactions. Specifically, within the explosives industry chemical kinetics influences safety, stability and performance of explosives.

Safety & Stability : recent studies have shown the use of a kinetic model can be used to predict impact sensitivity [3] but most chemical kinetics surrounding safety is used around the shelf life and storage of the explosives [4] [5] [6] [7].

Performance: Most performance based kinetics are around activation energies for energetic materials, how and why they are influenced and how they could be improved [8] [9] [10].

2. Materials & Method

The materials and method used can be found in Appendix 1.

3. Observations

Data given and collected are as follows:

	<i>KI</i>	<i>Vitamin C</i>	<i>Starch</i>	<i>Water</i>	<i>H₂O₂</i>	
<i>unit?</i>	Molarity	0.20	0.010	1%	-	(3%) 0.891
<i>mL used in Trial 1:</i>	3	1.5	1	18.5	1	
	3	1.5	1	17.5	2	
	3	1.5	1	16.5	3	
	3	1.5	1	15.5	4	
<i>mL used in Trial 2:</i>	1	1.5	1	18.5	3	
	2	1.5	1	17.5	3	
	3	1.5	1	16.5	3	
	4	1.5	1	15.5	3	

Table 1: Data given and collected for the experiment.

Molarity of H₂O₂:

$$\text{Vol} = \frac{\text{Mass}}{\text{Density}} \times \frac{100}{1.01} = 99.01 \text{ ml}$$

$$\text{Mole} = \frac{\text{Mass}}{\text{Mr}} \times \frac{3\text{g}}{34\text{g mol}^{-1}} = 0.0882 \text{ mole}$$

$$\text{Molarity} = \frac{\text{Moles of H}_2\text{O}_2}{\text{Volume of solution (L)}} = \frac{0.0882}{0.09901} = 0.891 \text{ M}_{\text{correct}}$$

To find the final molarity of the solution the following calculations were performed:

$$M_{\text{Final}} (\text{mol/L}) = \frac{L \text{ of H}_2\text{O}_2}{0.025 \text{ Total reaction solution L}} \times 0.891 (M)$$

$$M_{\text{Final}} (\text{mol/L}) = \frac{L \text{ of KI}}{0.025 \text{ Total reaction solution L}} \times 0.20 (M)$$

sig figs

[H₂O₂]₀	t	Log ([H₂O₂]₀)	1/t	<i>log(1/t)</i>	[I⁻]₀	t	Log ([I⁻]₀)	1/t	<i>log(1/t)</i>
0.036	200	-1.443697499	0.005		0.008	175	-2.243038049	0.005714	
0.071	100	-1.148741651	0.01		0.016	108	-2.033423755	0.009259	
0.11	80	-0.958607315	0.0125		0.024	69	-1.838849091	0.014493	
0.14	66	-0.853871964	0.015151		0.032	36	-1.556302501	0.027778	

Table 2 Values for Trial 1: [H₂O₂]₀, Time, Log ([H₂O₂]₀) and 1/t.

Table 3 Values for Trial 2: [I⁻]₀, Time, Log ([I⁻]₀) and 1/t.

Rate constant equation:

$$k = \frac{1/t}{[H_2O_2]^m [I^-]^n}$$

$$\text{Trial 1 } k: \frac{0.005}{[0.036]^1 \times [0.20]^2} = 3.47$$

4. Results

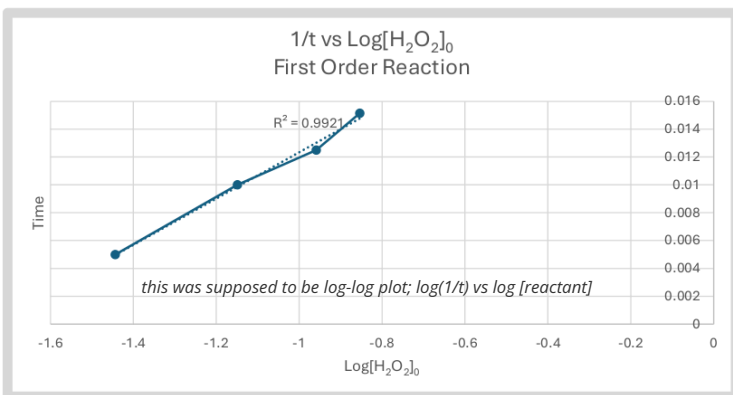


Figure 2: Graph depicting $1/t$ vs $\text{Log}[\text{H}_2\text{O}_2]_0$

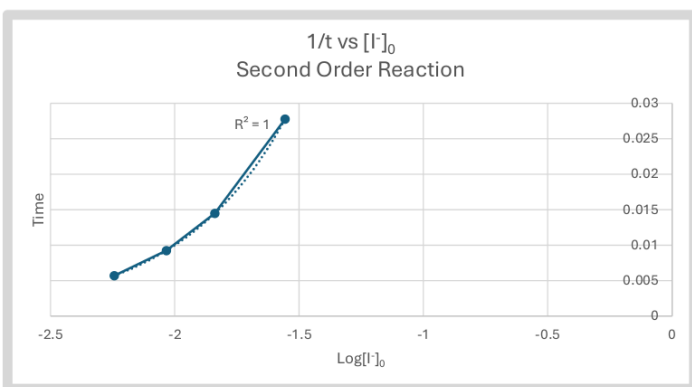


Figure 3: Graph depicting $1/t$ vs $[\text{I}^-]_0$

5. Discussion

For trial 1, table 2 shows that between experiments 1 and 2, doubling the concentration roughly doubles the initial reaction rate shown in $1/t$, this initially would reflect that the reaction rate is proportional, however when comparing experiments 1 and 4, quadrupling the concentration shows a discrepancy of 0.31 this could be down to a variety of factors, more than likely due to accumulated rounding errors that stem from

slight discrepancies in volume added due to any remaining liquid clinging to the initial vessel and not being added to the final solution. This reaction rate proportionality is characteristic of a first order reaction [11] this is also shown in the graph in figure 2, due to this the rate constant can be calculated using any row in the table resulting in $K=3.47$ for trial 1.

For trial 2, table 3 shows that between experiments 1 and 2, doubling the concentration roughly doubles the initial reaction rate shown in $1/t$, this again would reflect that the reaction rate is proportional [11] there is a discrepancy again of 0.002169 and because as mentioned before due to rounding errors moving forwards, however this proportionality is again characteristic of a first order reaction, however when used in conjunction for the whole experiment would put the overall order at second order [12]. This second order can be seen in the graph data from figure 3, the slope of the curve is more characteristic of a second order reaction *to determine the orders, m and n , you were supposed to use the slope of the trend-line*

6. Conclusion

This experiment proved the initial point, but moving forwards there should have been more experiments per concentration to get an average time. Next time the addition of rerunning the experiment again at an increased temperature. It would be expected that the reaction time should decrease with increased temperature and due to this faster reaction rate, a higher rate constant would be observed. The rate is dependent on activation energy, using both sets of measurements at $Temp_1$ and $Temp_2$ could then subsequently be used to calculate activation energy as per the Arrhenius equation [13].

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Appendix 1.

Materials & Method.

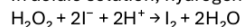
ACDA 6020 LAB SESSION #2

RATE OF REACTION (Advanced)

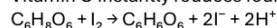
AIM OF THE EXPERIMENT: To determine the rate law for the formation of iodine in the reaction between hydrogen peroxide and iodide ions in aqueous solution, using vitamin C (ascorbic acid) as a reducing agent and starch as an indicator.

INTRODUCTION

In acidic solution, hydrogen peroxide oxidises iodide ions to iodine:



Vitamin C instantly reduces iodine back to iodide:



As long as vitamin C is present, no blue colour appears. When it is consumed, iodine accumulates and forms a blue starch–iodine complex. The time until this colour appears is the clock time (t), which reflects the rate of iodine formation.

In [Part 1](#) you will determine the order of the reaction in $[\text{H}_2\text{O}_2]$ and in [Part 2](#) – order of reaction in $[\text{I}^-]$.

Why we use the “clock” method

In most kinetics experiments, the rate of reaction is found by measuring how the concentration of a reactant or product changes continuously with time. In this iodine clock experiment, however, we cannot easily monitor the gradual formation of iodine because it is immediately reduced by vitamin C until that reagent is fully consumed.

To overcome this, we use a clock approach: the moment the blue starch-iodine colour appears corresponds to the instant when all vitamin C has been used up and a small, fixed amount of iodine suddenly accumulates.

That fixed iodine amount is the same in every trial, so the time taken to reach it (the clock time, t) depends only on how fast iodine is being produced.

If iodine forms quickly, the blue colour appears sooner (small t); if formation is slower, the colour takes longer (large t).

This means the average rate of the reaction is inversely proportional to the measured clock time:

$$\text{Rate} \propto 1/t$$

By varying one reactant concentration at a time and comparing the reciprocal of the times, we can determine how the rate depends on each concentration and hence deduce the rate law for the reaction.

PROCEDURE

Put on the lab coat, safety goggles, and gloves. Make sure that long hair is tied.

IMPORTANT: Record all your observation in the data sheets at the end of this script.

Lab tools and supplies:

- 50 mL conical flasks x4
- Glass pipettes x2 (*pipette 1 used for KI and ascorbic acid; pipette 2 – H_2O_2*) and pipette filler
- Graduated cylinder or small beaker and test tube
- Stopwatch
- Glass stirring rod

Chemicals:

Solution A (in conical flask):

- 0.20 M KI solution
- 0.010 – 0.025 M ascorbic acid (Vitamin C) solution
- 1 % starch indicator
- Distilled water

Solution B (in test tube):

- 3% H_2O_2 solution

Procedure:**I. Order of the reaction in [H_2O_2]**

1. In four conical flasks prepare solution A according to the table. Use one glass pipette to measure the required volumes of KI and Vit C solutions. Add ~1 mL of starch solution.
2. With the second glass pipette measure the required volume of H_2O_2 (solution B) and transfer it in the test tube.
3. Pour quickly solution B from the test tube to solution A in the conical flask, stir and start the timer as soon as the solutions are mixed.
4. Record the time until the first permanent blue colour appears.

Trial	A: KI 0.20 M (mL)	A: Vit C 0.010 M (mL)	A: Starch 1% (mL)	A: Water (mL)	B: H_2O_2 3% (mL)
I.1	3.0	1.5	1.0	18.5	1.0
I.2	3.0	1.5	1.0	17.5	2.0
I.3	3.0	1.5	1.0	16.5	3.0
I.4	3.0	1.5	1.0	15.5	4.0

II. Order of the reaction in [I^-]

5. In four conical flasks prepare solution A according to the table. Use one glass pipette to measure the required volumes of KI and Vit C solutions. Add ~1 mL of starch solution.
6. With the second glass pipette measure the required volume of H_2O_2 (solution B) and transfer it in the test tube.
7. Pour quickly solution B from the test tube to solution A in conical flask, stir and start the timer as soon as the solutions are mixed.
8. Record the time until the first permanent blue colour appears.

Trial	A: KI 0.20 M (mL)	A: Vit C 0.010 M (mL)	A: Starch 1% (mL)	A: Water (mL)	B: H_2O_2 3% (mL)
II.1	1.0	1.5	1.0	18.5	3.0
II.2	2.0	1.5	1.0	17.5	3.0
II.3	3.0	1.5	1.0	16.5	3.0
II.4	4.0	1.5	1.0	15.5	3.0

FINAL GRADE

42/100

GENERAL COMMENTS

Engagement with Literature 6/15

A long list of references was provided, but many were not relevant to the kinetic experiment. Although the list contained a mix of credible academic and non-academic sources, most supported only the broader importance of kinetics in OME rather than the kinetic theory required for this study. Literature was not used to explain the rate law, reaction orders, or the iodine clock mechanism. The clock method was referenced, but key chemical principles were not supported by citations. Several references contained incorrect author names, and IEEE formatting was inconsistent. Overall, engagement with literature was present but poorly aligned with the experiment.

Knowledge and Understanding 10/20

An attempt was made to explain the redox chemistry between H_2O_2 , I^- , I_2 , and vitamin C, and the role of vitamin C was partially understood. However, the redox chemistry was described inaccurately, with missing or incorrect equations and misunderstanding of the roles of vitamin C and starch. The clock reaction was not fully understood; no explanation was given for why the rate is proportional to $1/t$. Fundamental definitions (rate law, reaction order, rate constant) were not provided, and log-log analysis was not explained. The misconception that the rate constant changes at constant temperature indicates a major theoretical gap. The belief that a first-order rate plot of rate vs concentration should be linear is also incorrect. OME relevance was mentioned in terms of sensitivity, ageing, stability, and performance, but the discussion was not linked to the experimental chemistry or full lifecycle considerations.

-

Cognitive and Intellectual Skills 8/20

The student attempted the kinetic method but did not apply it correctly. Axes were not log-transformed, making the slopes meaningless for determining reaction orders. Orders were incorrectly identified, and no regression equations were used to extract m and n . Experimental repetition at different temperatures was mentioned, but the purpose related to activation energy was not explained. No systematic error analysis was provided. While a conceptual link between kinetic accuracy and OME performance was acknowledged, the reasoning was not well developed.

Practical Application Skills 12/30

Several calculation errors were present, and tables of clock times were incomplete. Units were missing, and significant figures were inconsistent. Graphs were incorrectly constructed: axes were reversed, log transforms were not applied, and best-fit line equations were not displayed. As a result, slopes were not used for determining reaction orders. Although the final values for m and n become close to 1 when recalculated correctly, this accuracy reflects the underlying method rather than the student's own calculations. The rate constant k was calculated following the handout, but because KI concentrations and plots were incorrect, no valid kinetic information could be extracted from the student's data.

Transferable Skills 6/15

The report was written in a clear formal tone, with consistent passive voice and appropriate terminology. Tables and figures were labelled and referenced, and overall presentation was neat. However, several figures were incorrectly labelled—axes were marked as logarithmic despite non-logged data. Some content appeared in the wrong sections, and grammatical issues were present. The structure did not fully support the scientific narrative.

-

PAGE 1



Comment 1

This is not really the abstract; it's rather the aim of the report. An abstract is a concise summary of a research paper or report, typically ranging from 150 to 300 words, that outlines the main objectives, methods, results, and conclusions of the study.

PAGE 2

PAGE 3

Text Comment. what does this indicator indicate?

PAGE 4

Text Comment. unit?

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PAGE 5

Text Comment. this was supposed to be log-log plot; $\log(1/t)$ vs $\log[\text{reactant}]$

Text Comment. this method can only be applied to the a single reaction for which change in concentration can be monitored; does not apply to a clock reaction

PAGE 6

Text Comment. to determine the orders, m and n, you were supposed to use the slope of the trend-line

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