

MF Lab2

by Max Fisher

Submission date: 06-Nov-2025 11:45AM (UTC+0000)

Submission ID: 266856910

File name: 32325_Max_Fisher_MF_Lab2_695600_1622951740.docx (58.93K)

Word count: 1913

Character count: 10394

ACDA6020 LAB #2

Rates of Reaction: Iodine Clock

Max Fisher

2113637

Contents

ACDA6020 LAB #2	1
Rates of Reaction: Iodine Clock	1
Table of figures.....	1
Table of tables.....	1
Introduction and Hypothesis	2
Importance of Kinetics in OME.....	2
Method.....	2
Results	3
Discussion.....	6
Conclusion:	7
References	7

Table of figures

Figure 1 H ₂ O ₂ Varied.....	4
Figure 2 KI Varied.....	4
Figure 3 KI Varied 2.2 excluded.....	5
Figure 4 KI Varied 2.1 excluded.....	5

Table of tables

Table 1 Reagent quantities.....	3
Table 2 results.....	3
Table 3 Constant K calculation	6
Table 4 Constant K calculation (excluding 2.1).....	6

Introduction and Hypothesis

Across modern life many different elements are influenced by the presence of varying rates of reaction. Examples such as heat packs [1], internal combustion [2] cars and especially energetics all have different effects based on the rate of the reactions present. Whilst there is a plethora of different influences on these, the rates of reaction can be calculated and predicted. This is essential for industrial processes, where the competing influences of demand cost and safety often collide. This experiment attempts to characterise the rate of reaction for the common 'iodine clock', this reaction clearly demonstrates when the reactants have been consumed through an obvious colour change: the liquids present change from colourless to a deep blue. The following report will detail the authors findings and provide a rate calculation. It is expected that the H_2O_2 will have a ~~have a~~ similar effect on the rate of reaction to KI this is because whilst H_2O_2 is a strong oxidiser, KI is also considered a strong reducer. It is expected that both 'sub' reactions will of the first order.*what is this hypothesis based on? no supporting literature*

Reaction 1: liberation of Iodine from potassium iodide $\text{H}_2\text{O}_2 + 2\text{KI} + 2\text{H}^+ \rightarrow \text{I}_2 + 2\text{H}_2\text{O} + 2\text{K}^+$ [3]

Reaction 2: Iodine reacts with Vitamin C to form Potassium iodide and the oxidised form of ascorbic acid $\text{C}_6\text{H}_6\text{O}_6 + \text{I}_2 + 2\text{K}^+ \rightarrow \text{C}_6\text{H}_6\text{O}_5 + 2\text{H}^+ + 2\text{KI}$

Reactions 1 and 2 continue back and forth until all vitamin C is oxidised.

'Reaction' 3: when the Vitamin C is used, excess I_2 forms, which then mixes with starch to form the blue colour. This is not a chemical reaction but form of mixture $(n\text{C}_6\text{H}_{10}\text{O}_5 + m\text{I}_2 \rightarrow [\text{C}_6\text{H}_{10}\text{O}_5 + m\text{I}_2] \text{ complex})$ ^[Ref]

Importance of Kinetics in OME

Kinetics is of great importance to Ordnance munitions and explosives. Through varying rates of reactions is how many energetic effects are created; these can range from the near instantaneous reactions in military explosives, which propagate above the speed of sound (of the material), to safety fuze compositions, which burn slowly and repeatably. The difference in these effects is only possible due to having varied and carefully curated reaction kinetics. By understanding and modelling [4] the effects of kinetics on the decompositions it enable safer operations both during the production cycle and during the through life surveillance [5]. The majority of energetics degrade over time, and through understanding the factors that affect these mechanisms both performance and safety can be improved.

Method

The experiment was conducted to the below method:

Reagents A (KI, Vitamin C, Starch and Water), were added single conical flask. These were measured using a pipette dropper.

Reagent B (H_2O_2) was added to a small test tube using a pipette dropper.

See Table 1 for quantities.

(only two droppers were used, but these were cleaned thoroughly using distilled water between measuring reagents).

Reagent B was then added to A in the conical flask, and a timer started. Each sample was stirred manually for 25s and the time of colour change recorded.

Control variables:

Quantities of starch and vitamin C, (trial 1: KI), (trial 2 H₂O₂). Time stirred

Independent variables

Quantity of **Water**, (trial 1 H₂O₂), (trial 2 KI)

Dependent variable:

Time taken for colour change.

Table 1 Reagent quantities

Trial	A: KI 0.2M (ml)	A: Vit C 0.010M (ml)	A: Starch 1%(ml)	A: Water (ml)	B: H ₂ O ₂ 3% (ml)
1.1	3.0	1.5	1.0	18.5	1.0
1.2	3.0	1.5	1.0	17.5	2.0
1.3	3.0	1.5	1.0	16.5	3.0
1.4	3.0	1.5	1.0	15.5	4.0
2.1	1.0	1.5	1.0	18.5	3.0
2.2	2.0	1.5	1.0	17.5	3.0
2.3	3.0	1.5	1.0	16.5	3.0
2.4	4.0	1.5	1.0	15.5	3.0

Results

The results generated were as follows:

Reactant molar mass H₂O₂ = 2*1+ 2*16 = 34g/mol

H₂O₂ 3% solution therefore 30g/1000ml so 30/34g = 0.88235mol/L =
(volume*(0.88235/1000))/0.025

KI = 0.2m/L = 0.0002m/ml - 25ml all in so overall concentration of KI = amount of
KI mol*0.0002/0.25

Table 2 results

	Trial	Volume of reactant (ml)	[Reactant] ₀ (M)	Clock Time (s)	1/t (s ⁻¹)	Log [Reactant]	Log(1/t)
H ₂ O ₂ varied	1.1	1.0	0.035	110	1/110	-1.452	-2.041
	1.2	2.0	0.071	60	1/60	-1.151	-1.778
	1.3	3.0	0.106	53	1/53	-0.975	-1.724
	1.4	4.0	0.141	47	1/47	-0.850	-1.672
KI varied	2.1	1.0	0.008	65	1/65	-2.097	-1.813
	2.2	2.0	0.016	74	1/74	-1.796	-1.869
	2.3	3.0	0.024	57	1/57	-1.620	-1.756
	2.4	4.0	0.032	49	1/49	-1.495	-1.690

Using the experimental data from Table 1, Figures 1-3 show the data plots for both the H_2O_2 trials and the KI trials. From figure 1, the order of the reaction is clearly below a first order. This is evident from the equation shown ($y=0.612x-1.1263$), the increase in Hydrogen peroxide clearly increases the rate of the reaction, but from the experimental data is below the first order (0.61). The R^2 value shows low deviation from the trendline calculated within excel, adding confidence to this assessment of the experimental data.

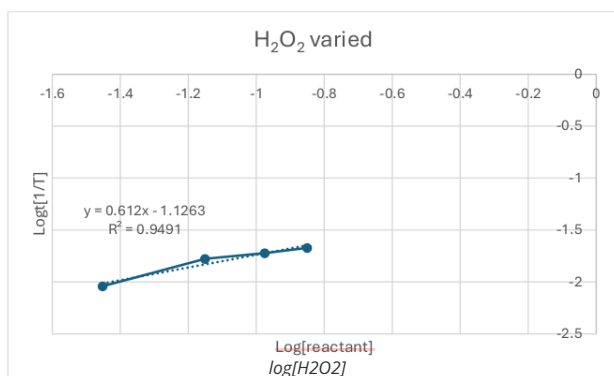


Figure 1 H₂O₂ Varied

Figure 2 shows the KI variation, this equation shows a much different relationship between the concentration and time, in this case the equation is $y=0.2035x-1.4255$, this is again sub-first order, but significantly lower than in figure 1. However, when looking at the R^2 value this shows a poor correlation between the expected values and the values received. This appears to be due to the result at 2.2, where the experimental value is significantly below the trendline or 2.1 where it is clearly not in line with the other three point.

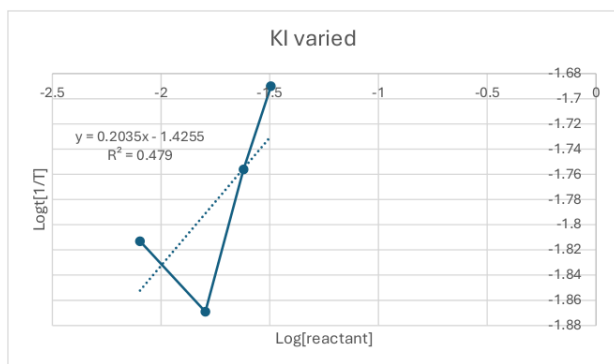


Figure 2 KI Varied

Figure 3 shows the same graph as Figure 2, but with the results of sample 2.2 excluded. The

equation from this is $y=0.1808x-1.4389$ with an R^2 value of 0.8707. This means that the trendline is much closer to the expected values. Whilst it may appear that this is an anomaly, due to a lack of data, this cannot be discounted. As such trial 2.2 cannot be excluded from the overall rate calculation.

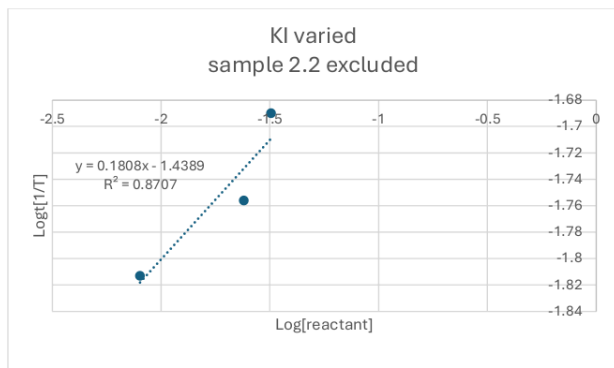


Figure 3 KI Varied 2.2 excluded

Figure 4 shows the trendline with sample 2.1 excluded, this gradient is much higher than the others and much closer to the trendline in experiment 1. In addition, the R^2 value of 0.9972 suggests that the trendline is extremely well correlated with the data points.

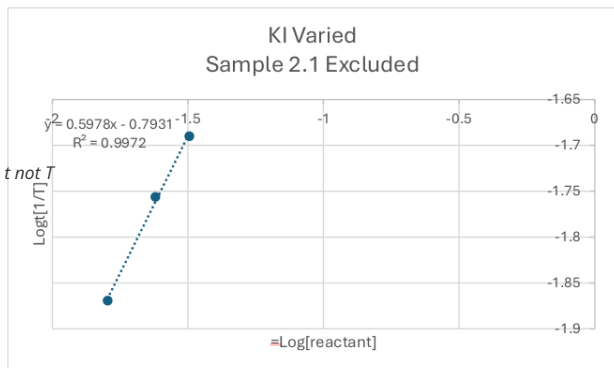


Figure 4 KI Varied 2.1 excluded

Overall rate law can be calculated by using the following equation

Rate= $K \cdot [H_2O_2]^m \cdot [KI]^n$ where K is a constant m is the H2O2 logs gradient, and n is the KI logs gradient.

M=0.612 n=0.2035.

to calculate K, the author took an average of trials 1.1, 1.4, 2.1 and 2.4. samples

using the following equation

$$K = (1/t) / ([H_2O_2]^m \cdot [KI]^n)$$

Table 3 Constant K calculation

trial	1/t	H2O2	KI	Value for K	missing units
1.1	1/110	0.035	0.024	0.151	
1.2	1/60	0.071	0.024	0.179	
1.3	1/53	0.106	0.024	0.159	
1.4	1/47	0.141	0.024	0.151	
2.1	1/65	0.106	0.008	0.162	
2.2	1/74	0.106	0.016	0.124	
2.3	1/57	0.106	0.024	0.148	
2.4	1/49	0.106	0.032	0.162	
			average	0.1545	

This provided the overall rate:

$$\text{Rate law} = 0.155 \cdot [H_2O_2]^{0.612} \cdot [KI]^{0.2035}$$

Table 4 Constant K calculation (excluding 2.1)point 2.1 should be excluded

trial	1/t	H2O2	KI	Value for K
1.1	1/110	0.035	0.024	0.658
1.2	1/60	0.071	0.024	0.783
1.3	1/53	0.106	0.024	0.693
1.4	1/47	0.141	0.024	0.656
2.2	1/74	0.106	0.016	0.633
2.3	1/57	0.106	0.024	0.645
2.4	1/49	0.106	0.032	0.631
			average	0.6712

$$\text{Rate law} = 0.671 \cdot [H_2O_2]^{0.612} \cdot [KI]^{0.598}$$

The overall calculated order is 0.8155. This is significantly below the expected order. However, if value 2.1 is ignored from the KI calculation the overall order rises to an overall order of 1.21.

Discussion

The rate generated experimentally for this reaction is as follows:

$$\text{Rate law} = 0.155 \cdot [H_2O_2]^{0.612} \cdot [KI]^{0.2035} \text{ (or if 2.1 excluded) } = 0.671 \cdot [H_2O_2]^{0.612} \cdot [KI]^{0.598}$$

This Shows both reactions as sub-first order, this is substantially different to literature sources where both the variation of Hydrogen peroxide [6]and the variation of the Potassium Iodide [7] are considered as full first order reactions. [8] the Experimental results show the concentration of Hydrogen Peroxide had a greater effect on the rate of reaction than that of the Potassium iodide, the reasons for this are unclear, but may be due to experimental error. If experiment 2.1 is discounted, the iodine clock reaction is affected similarly by both H₂O₂ and KI. Whilst the results could be rounded to produce a first order reaction, the author feels that this would not represent the results most faithfully but would produce results similar to the literature.

To improve the results, it would be recommended to repeat the experiment multiple times, this will reduce the effect of random variation or experimental error. It is believed by the author that trial 2.1 is an anomalous result, but due to the limited sample size, this cannot be proven, and as such is potentially poisoning the results, reducing the order significantly. Another element that was not controlled was the temperature of the reaction and reaction vessels. It is also noted that the indicator was added using a pipette dropped, instead of the calibrated instruments, whilst this will not cause a drastic change as this was the indicator solution, the variation in this could have caused an increase (or decrease) in the relative concentrations which could again have affected the results significantly. In addition, the sharing of equipment between reactants could have introduced significant levels of contamination to the experimental results. Despite thorough cleaning between serials, it is normally expected to use one measuring device per reactant, eliminating the risk of contamination between serials.

Conclusion:

Overall, this experiment was useful to examine the overall rates of reaction and ways in which there are affected by the reactants. The experimental results are a poor reflection of the literature results but could be caused by a number of experimental factors. These would likely be resolved in repeated trials and through more careful control of external factors such as temperature and use of laboratory equipment such as magnetic stirrers and individual pipettes for each reactant.

References

[D. Whyte, "Bring on the Heat! Investigating Exothermic Reaction Rates," Science Buddies, 2011 11 2020. [Online]. Available: https://www.sciencebuddies.org/science-fair-projects/project-ideas/Chem_p054/chemistry/investigating-exothermic-reaction-rates. [Accessed 6 11 25].

[S. Bragg, "Application of Reaction Rate theory to combustion chamber analysis," Thornton 2 research centre, 1953.
]

[L. T. -. M. T. Saraswathiamma., "Kinetics- Iodine Clock using Vitamin C and the Calculation of 3 Initial Rate-Home," [Online]. Available:
] https://chem.libretexts.org/Courses/Minnesota_State_Community_and_Technical_College/CHEM_1112%3A_General_Inorganic_Chemistry_II_Lab_Manual-Online_Section/02%3A_Experiments/2.08%3A_Kinetics-_Iodine_Clock_using_Vitamin_C_and_the_Calculation_of_Initial_Rate-. [Accessed 5 11 25].

[W. M. H. L. E. F. C. Souers, "Kinetic Modeling of Non-Ideal Explosives with CHEETAH," in 4 *International Detonation Symposium*, Snowmass, 1998.
]

[L. C. B. P. G. Jordan Ehrhardt Julien Glorian, "Detailed kinetic mechanism for nitrocellulose
5 low temperature decomposition," *Combustion and FLame*, vol. 2, no. 258, 2023.

]

[Libre Chemistry Texts, "Iodine Clock Reaction," [Online]. Available:

6 https://chem.libretexts.org/Courses/Saint_Marys_College_Notre_Dame_IN/Chem_122L%3A

] _Principles_of_Chemistry_II_Laboratory_(Under_Construction___)/13%3A_Iodine_Clock_Rea
ction. [Accessed 26 10 2025].

[S. Hird, "Iodine Clock Reaction (AQA A Level Chemistry)," [Online]. Available:

7 <https://www.savemyexams.com/a-level/chemistry/aqa/17/revision-notes/8-advanced->

] practical-skills/8-1-physical-chemistry-practicals/8-1-5-iodine-clock-reaction/. [Accessed
26 10 2025].

[Royal Society of Chemistry, "Iodine clock reaction demonstration," 16 Jan 2018. [Online].

8 Available: <https://edu.rsc.org/experiments/iodine-clock-reaction-demonstration->

] method/744.article. [Accessed 26 10 2025].

63/100

Engagement with Literature 10/15

A large number of references was included, but many were not relevant to kinetic analysis or the iodine clock reaction. Some citations were attached to chemically incorrect statements, and redox chemistry was not accurately supported by the cited literature. A reasonable mix of academic and non-academic sources was used, though academic references mainly supported OME context rather than kinetic theory. Iodine-clock sources were not used to build the kinetic explanation. IEEE formatting contained several errors.

Knowledge and Understanding 12/20

The redox processes involving H_2O_2 , I^- , I_2 , and vitamin C were explained correctly, and the role of vitamin C was understood. However, the proportionality of rate to $1/t$ was not explained, and key definitions (rate law, reaction order, rate constant) were missing.

The hypothesis lacked support, and variables were identified with minor errors. The rationale for using log-log plots was not discussed. Links to OME sensitivity, ageing, stability, and performance were made, though without full lifecycle context.

Cognitive and Intellectual Skills 13/20

Experimental errors were acknowledged but not deeply analysed. Log-log plots were evaluated thoroughly, and reaction orders were extracted from slopes, though the overall hypothesis remained unsupported.

Interpretation showed some analytical ability, but underlying causes of deviations and limitations of the kinetic method were not fully examined.

Practical Application Skills 20/30

Calculations were correct, and the table of times was complete. Units and significant figures were inconsistent. Log-log plots were mostly correct but had axis-labelling issues. Best-fit lines were used appropriately to determine reaction orders.

The rate constant k was calculated following the handout, though clearer numerical presentation and improved precision were needed.

Transferable Skills 8/15

A formal, objective tone was maintained, with consistent passive voice and correct scientific terminology. Tables and figures were labelled and referenced. Minor grammar errors and some formatting weaknesses were present.

PAGE 1

PAGE 2

Strikethrough.

Text Comment. what is this hypothesis based on? no supporting literature

Text Comment. [Ref]

PAGE 3

Strikethrough.

Text Comment. sig figs

PAGE 4

Strikethrough.

Text Comment. $\log[\text{H}_2\text{O}_2]$

PAGE 5

Text Comment. t not T

Strikethrough.

PAGE 6

Text Comment. missing units

Text Comment. point 2.1 should be excluded

PAGE 7

Strikethrough.

PAGE 8

Strikethrough.