

LAB 2

by Libbie Bush

Submission date: 14-Nov-2025 02:50PM (UTC+0000)

Submission ID: 267678653

File name: 42747_Libbie_Bush_LAB_2_695604_2058802486.docx (77.6K)

Word count: 1345

Character count: 7005



LAB: RATE OF REACTION

L6 -Advanced Chemistry of OME (ACDA6020)

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Table of Contents

Table of Equations	2
Table of Tables	2
Aim	2
Variables	2
Theory	3
Results, Observation and Discussion	4
References (IEEE)	Error! Bookmark not defined. ???
Appendix – Method and Equipment	8

Table of Equations

Equation 1 = Rate	3
Equation 2 - Rate order	3
Equation 3 - Rate equation for reaction	3
Equation 4 - Value of K	6

Table of Tables

Table 1 - Part one	4
Table 2- Part 2	4
Table 3 - results	5
Table 4 - Rate Graph	6

Aim

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. *clear aim*

Variables

Independent – concentration of H_2O_2 and I^-

Dependant – time till blue colour appears

Control – Volume of starch, vitamin C concentration, mixing procedure

Theory

The rate of reaction is defined by how long it takes for the reactants to convert to products. (1)The rate of a reaction can be affected by multiple things, such as temperature, concentration and catalysis. A formula to help determine the rate of reaction is as follows [2]

$$\text{Rate} = -\frac{d[\text{Reactant}]}{dt} = \frac{d[\text{Products}]}{dt} \text{ Equation 1 = Rate}$$

A reaction order shows how much that concentration affects the rate of reaction, the equation for this is shown below

$$\text{Rate} = k[A]^m[B]^n \text{ Equation 2 - Rate order}$$

k K = rate constant

$[A]$ and $[B]$ = the concentration of reactants

m and n = orders of reactants

An order of a reaction shows how much the rate depends on concentration. The overall order of the reaction is $m+n$. If the reaction is 0 order, then the rate does not depend on concentration. If the reaction is 1st order, then the rate is directly proportional to the concentration. If the reaction is 2nd order, then the rate depends on the square of the reactant's concentration. [3] *very good, concise definition*

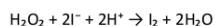
Iodine clock

The iodine clock experiment is a way to measure the rate of a reaction. The formula for this is shown below

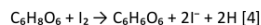
$$\text{Equation 3 - Rate equation for reaction}$$

$$\text{Rate} = k[\text{H}_2\text{O}_2]^m[\text{I}^-]^n$$

When the hydrogen peroxide is added, it goes through the following reaction, which produces iodine



However, this iodine instantly reacts with the vitamin c to produce iodide in the following reaction



Once all the vitamin C has been used up, then the iodine is produced, which then reacts with the starch to make a blue-coloured solution. This marks the end of the reaction.

The following amounts were used in the reaction

Table 1 - Part one

Trial	A: KI 0.20 M (mL)	A: Vit C 0.010 M (mL)	A: Starch 1% (mL)	A: Water (mL)	B: H ₂ O 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

Table 2- Part 2

Trial	A: KI 0.20 M (mL)	A: Vit C 0.010 M (mL)	A: Starch 1% (mL)	A: Water (mL)	B: H ₂ O 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

In OME

In the OME industry, this is used to ensure that the chemical compounds react at the expected rate. Too low could cause a hang or misfire within a firing system. Too fast could be a safety hazard [5]

Results, Observation and Discussion

Calculate the molarity of Hydrogen peroxide after dilution:

Molarity=mole/volume

Moles before dilution and after dilution are equal

Mole=molarity x volume

molarity x volume (before) = molarity x volume (after)

Molarity after = [molarity x volume (before)]/volume (after)

Molarity before =

Assuming 100g of solution, then 3% Hydrogen peroxide 3g

Moles = mass/molar mass

$$\text{Molar mass} = 1 + 1 + 16 + 16 = 34$$

$$\text{Moles} = 3/34 = 0.08823 \text{ mol}$$

$$\text{Molarity} = \text{moles/volume}$$

units are missing

$$\text{Molarity} = 0.08823/0.1 \text{ mL}$$

$$\text{Molarity} = 0.08823 \text{ mol/L}$$

Volume before is the volume of hydrogen peroxide added to the solution

Volume after is the overall solution shown in Tables 1 and 2

Using this information in tables 1 and 2, the concentration of reactants is shown below:

Calculate the molarity of Iodine after dilution:

$$\text{Molarity} = \text{mole/volume}$$

Moles before dilution and after dilution are equal

$$\text{Mole} = \text{molarity} \times \text{volume}$$

$$\text{molarity} \times \text{volume (before)} = \text{molarity} \times \text{volume (after)}$$

$$\text{Molarity after} = [\text{molarity} \times \text{volume (before)}] / \text{volume (after)}$$

$$\text{Molarity before} = 0.2 \text{ (from tables 1 and 2)}$$

Volume before is the volume of hydrogen peroxide added to the solution

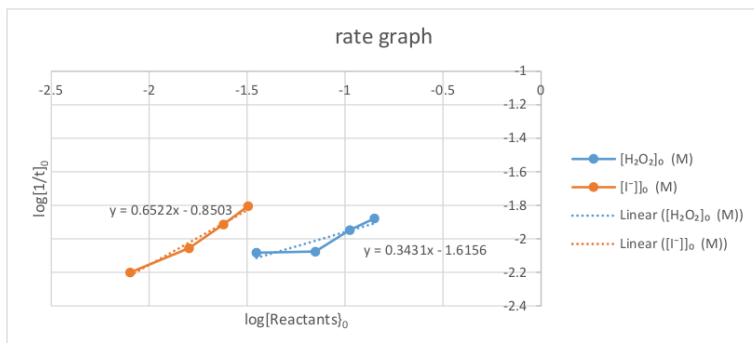
Volume after is the overall solution shown in Tables 1 and 2

Using this information in tables 1 and 2, the concentration of reactants is shown below:

Table 3 - results

Reaction	Time (s)	1/t	$\log[1/t]_0$	$[\text{H}_2\text{O}_2]_0$ (M)	$\log[\text{H}_2\text{O}_2]_0$	$[\text{I}^-]_0$ (M)	$\log[\text{I}^-]_0$
I.1	120.79	0.008279	-2.082031	0.035292	-1.45232373	0.024	-1.619788758
I.2	118.97	0.008405	-2.075437	0.070584	-1.15129373	0.024	-1.619788758
I.3	88.53	0.011296	-1.94709	0.105876	-0.97520247	0.024	-1.619788758
I.4	75.38	0.013266	-1.877256	0.141168	-0.85026374	0.024	-1.619788758
II.1	158.42	0.006312	-2.19981	0.105876	-0.97520247	0.008	-2.096910013
II.2	113.6	0.008803	-2.055378	0.105876	-0.97520247	0.016	-1.795880017
II.3	81.81	0.012223	-1.912806	0.105876	-0.97520247	0.024	-1.619788758
II.4	63.62	0.015718	-1.803594	0.105876	-0.97520247	0.032	-1.494850022

Table 4 - Rate Graph



From the graph, you can deduce the orders of the reactants (m and n) as these are the slope of the graph

$$n = 0.6522$$

$$m = 0.3431$$

Now that the orders are known, the value for the rate constant K can be calculated with the equation

Equation 4 - Value of K

$$k = (1/t) / ([H_2O_2]^m [I^-]^n)$$

The values of K are shown below in table 5

Table 5 - K values

Reaction	1/t	[H ₂ O ₂] ₀ (M)	[I ⁻] ₀ (M)	K
I.1	0.008279	0.035292	0.024	0.2969529
I.2	0.008405	0.070584	0.024	0.2376827
I.3	0.011296	0.105876	0.024	0.2779251
I.4	0.013266	0.141168	0.024	0.2957301
II.1	0.006312	0.105876	0.008	0.3179704
II.2	0.008803	0.105876	0.016	0.2821542
II.3	0.012223	0.105876	0.024	0.3007543
II.4	0.015718	0.105876	0.032	0.3205823

The values for k slowly change for each reaction; this value should stay constant even when the concentration changes. This shows that there is an external factor that influences this experiment, such as temperature *excellent observation!*

This experiment also relies on a manual timer and a visual colour change; this means that the individual's reaction time plays a big part in the result, and the colour change point could be seen as objective, as it would be almost impossible to get the exact point every time

Conclusion

This experiment gives you the reactants' order and values for k , however, the potential points of error in this experiment could decrease the credibility of the results. However, it could still offer a rough idea of what the individual may be looking at.

References

[1]

BBC, "Rates of reaction," *BBC Bitesize*, 2019. <https://www.bbc.co.uk/bitesize/guides/z3nbqhv/revision/1>

[2]

BBC Bitesize, "Rates of Reaction - Revision 2 - National 5 Chemistry - BBC Bitesize," *BBC Bitesize*, 2019. <https://www.bbc.co.uk/bitesize/guides/zct4fcw/revision/2>

[3]

S. Hird, "Iodine Clock Experiment - A Level Chemistry Revision Notes," *Save My Exams*, Oct. 26, 2024. <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>

[4]

LibreTexts, "13: Iodine Clock Reaction," *Chemistry LibreTexts*, Jan. 13, 2020. [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_Reaction](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_Reaction)

[5]

D. Oestreich, Ludger Lautenschütz, U. Arnold, and J. Sauer, "Reaction kinetics and equilibrium parameters for the production of oxymethylene dimethyl ethers (OME) from methanol and formaldehyde," *Chemical Engineering Science*, vol. 163, pp. 92–104, Jan. 2017, doi: <https://doi.org/10.1016/j.ces.2016.12.037>.

access dates missing for all sources

Appendix – Method and Equipment

Equipment

0.20 M KI solution

0.010–0.025 M Vitamin C (ascorbic acid) solution

1% starch indicator

3% H_2O_2 solution

Distilled water

50 mL conical flasks

Glass pipettes

small beaker

Test tube

Stopwatch

Method

Part 1: Vary H_2O_2 volume (1–4 mL) to determine order in $[\text{H}_2\text{O}_2]$.

Part 2: Vary KI volume (1–4 mL) to determine order in $[\text{I}^-]$.

1. Prepare Solution A in conical flasks: KI, Vitamin C, starch, and water according to the table.
2. Prepare Solution B in a test tube: H_2O_2 (volume varies by trial).
3. Quickly pour Solution B into Solution A, stir, and start the stopwatch immediately.
4. Record the time until the first permanent blue colour appears (clock time).
5. Repeat for all trials:

53/100

Engagement with Literature 7/15

Literature engagement was limited. Most sources were non-academic and used to support general kinetic concepts. One peer-reviewed OME paper was listed but not integrated into the discussion. Access dates were missing, and IEEE formatting was incomplete. Key claims such as the proportionality of rate to $1/t$, expected reaction orders, the vitamin C regeneration mechanism, and interpretation of trendlines, were not supported by literature. OME references were cited only in the list, not meaningfully used in the body of the report.

Knowledge and Understanding 12/20

The aim and variables were correctly stated. Reaction rate and reaction orders were defined, and the redox chemistry was described clearly. OME was mentioned briefly, but no lifecycle-based kinetic discussion was provided. However, the log-log plotting method was applied without explanation, and the rate constant k was used without definition or discussion of its meaning or units. The clock method was not mechanistically explained, including why $1/t$ reflects rate or how vitamin C delays iodine formation. The hypothesis was not supported with literature.

Cognitive and Intellectual Skills 10/20

Changes in the calculated rate constant were recognised as incorrect, and potential sources of error were identified, though not explored in depth. Graph linearity and R^2 values were not evaluated, and anomalous points were not discussed. The hypothesis was not critically assessed, and no justification was provided for treating the slopes as exactly first order ($m \approx 1$, $n \approx 1$). No discussion of measurement uncertainty, subjective colour detection, temperature variation, or mixing delays was included.

Practical Application Skills 16/30

Calculations were mostly correct, though units were missing and significant figures were inconsistent. Log-log plots were constructed correctly. However, the rate constant calculation was conceptually incorrect: the expression $k = (1/t)/([H_2O_2]^m[I^-]^n)$ was used as though $1/t$ were the true rate. Units of k were not derived. The method appendix was referenced in the text but not provided. Table formatting was inconsistent. Anomalous data points

were retained despite reducing linearity. The rate constant was calculated according to the handout, but the underlying conceptual errors reduced validity.

Transferable Skills 8/15

The report was structured appropriately, though clearer figure and table titles would improve readability. Occasional first-person writing reduced formality. Formatting errors were present, including a broken Table of Contents link and inconsistent equation formatting. Captions did not align with IEEE style. Overall clarity was acceptable but required refinement.

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