

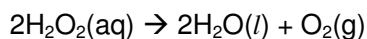
Name: () CT Group: Date:

Raffles Institution
Year 6 H2 Chemistry Practical 2014

Experiment 12: Reaction Kinetics

Worksheet Title: To determine the order of reaction with respect to H₂O₂ using a sampling method.

In this experiment, hydrogen peroxide decomposes catalytically according to the following equation:



The amount of hydrogen peroxide remaining in the mixture at any stage of the reaction may be determined by titration with aqueous potassium manganate(VII).

The order of reaction with respect to hydrogen peroxide and hence the rate equation and the rate constant for the reaction can be determined by monitoring the amount of hydrogen peroxide left at various time intervals.

The following chemicals are provided:

FA1 is aqueous hydrogen peroxide

FA2 is 0.0200 mol dm⁻³ aqueous potassium manganate(VII)

FA3 is aqueous sodium hydroxide

FA4 is aqueous iron(III) chloride

Dilute sulfuric acid

The following apparatus are provided:

Stopwatch

50 cm³ measuring cylinders

50 cm³ burette

10 cm³ pipette

25 cm³ pipette

Procedure (Experiment 1)

1. Using a pipette, transfer 50.0 cm³ of **FA1** into a 250 cm³ beaker and to this, add 25 cm³ of **FA3** from a measuring cylinder.
2. Place 50 cm³ of **FA4** into a second measuring cylinder.
3. At a convenient time, add **FA4** to the beaker. Start a stop watch at the instant of mixing and agitate the contents of the beaker to ensure even mixing. (When mixed, **FA3** and **FA4** react together to produce a catalyst and decomposition of the hydrogen peroxide then begins.)
4. As soon as possible, use a pipette to transfer 25.0 cm³ of the mixture to a conical flask. At exactly **two** minutes from the start of the reaction, add rapidly about 20 cm³ of dilute sulfuric acid into the conical flask.
5. Take further 25.0 cm³ samples from the reaction mixture in the beaker and add dilute sulfuric acid, as above, at **six**, **nine** and **fourteen** minutes, respectively, after the initial mixing.
6. Titrate the hydrogen peroxide present in each of the four samples thus obtained against **FA2**. Record your results in the table on page 2.

Note: Only one accurate titration is possible for each sample.

Results

Time / minutes	2	6	9	14	0
Final burette reading / cm ³					
Initial burette reading / cm ³					
Volume of FA2 used / cm ³	28.80	18.80	14.50	10.20	36.30

Procedure (Experiment 2)

7. Pipette 10.0 cm³ of **FA1** into a conical flask, add an approximately equal volume of dilute sulfuric acid using a measuring cylinder and then titrate with **FA2** to find the volume required before the hydrogen peroxide has begun to decompose. Again, only one accurate titration is necessary. Record your results in the above table in the column headed '0' minute.

Note: Why is 10.0 cm³ of **FA1** used in Experiment 2?

Amt of **FA1** pipetted in Experiment 1 = Amt of **FA1** pipetted in Experiment 2

$$25.0/1000 \times 50.0/125 \times [\text{FA1}]_{\text{pure}} = V_{\text{FA1}}/1000 \times [\text{FA1}]_{\text{pure}}$$

$$V_{\text{FA1}} = 10 \text{ cm}^3$$

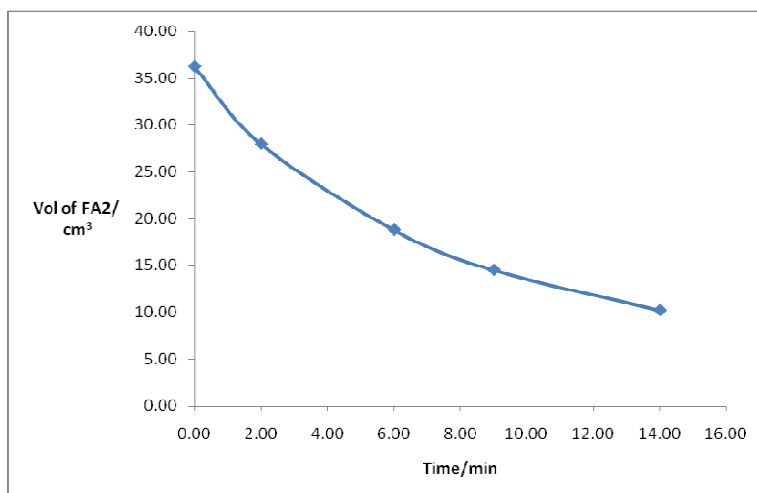
8. On a piece of graph paper, plot a graph of volume of **FA2** required for titration against time.

Answer the following questions.

- (a) Explain why about 20 cm³ of dilute sulfuric acid was rapidly added into the conical flask in **Step 4** of the procedure.

Sulfuric acid acts as a quenching reagent which stops the reaction. It reacts with the catalyst hence the reaction slows down drastically in the absence of a catalyst.

- (b) Use the graph of volume of **FA2** required for titration against time to determine the order of reaction with respect to hydrogen peroxide and hence the rate equation.



Vol of **FA2** required for titration is directly proportional to the amount of H_2O_2 remaining in the mixture, hence by plotting vol of **FA2** against time shows how H_2O_2 decreases with time.

Since the half life is constant, the order of reaction with respect to H_2O_2 is 1.

Rate equation: $\text{Rate} = k[\text{H}_2\text{O}_2]$

- (c) Determine the rate constant for the reaction.

Constant half-life = $\frac{1}{2}(7.1 + 7.4) = 7.25 \text{ min}$

Rate constant = $\ln 2 / 7.25 = 0.0956 \text{ min}^{-1}$

- (d) Discuss TWO main sources of error in both the procedure and measurements. Suggest improvements that could be made to the experimental procedure and measurements whilst following essentially the same general method. (You should consider modifications that will minimise errors and improve the reliability of the results.)

Sources of error	Improvements
Difficulty in ensuring the aliquots are quenched at the stipulated times	Use of a syringe to a quicker withdrawal/transfer of sampling solution. Ensure sample solutions are taken well before the quenching time Note: The use of a syringe might lead to a low precision where the volume of the solution is concerned.
Temperature of the reaction mixture may not be constant	Use a temperature-controlled water bath.
The accuracy of the titration particularly during the initial stage of the reaction (due to effect of human reaction time compared to short duration for the reaction)	Repeat the experiments to obtain average values. Or use a data logger.

- (e) Besides the method of withdrawing aliquots of the reaction mixture, quenching at specific times and followed by titrating, suggest another way by which the rate of reaction at various times can be obtained.

Connect the experimental set-up to a frictionless gas syringe. As the reaction proceeds, read the volume of oxygen collected at $\frac{1}{2}$ -minute intervals. Plot the graph of volume of oxygen evolved with time. Rate ($d[\text{O}_2]/dt$) at a particular time may be found by finding the gradient of the tangent of the graph at that time.