

Experiment

Catalytic decomposition of hydrogen peroxide (H_2O_2)



Prepared By:

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Date:

Room Temperature: 28 °C

Objective**To study the catalytic decomposition of hydrogen peroxide (H₂O₂).****Requirements**

(a) Glassware: Conical flask, Burette, Burette stand, Pipette, Measuring cylinder, Volumetric flask, Test tube, Watch glass, Beaker, Funnel, Stopwatch, Water bath.

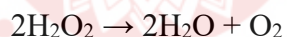
(b) Chemicals:

- 0.02 M KMnO₄
- 0.05 M FeCl₃+ 0.04 M HCl
- 1 % (v/v) H₂O₂ of Concentrated H₂O₂ [30 % w/v (or 100 volume)]
- Dilute H₂SO₄ (Approx. 5 N)

1. Principle

The rate of the chemical reaction depends on a number of factors, *e.g.* concentrations of the reactants, temperature, catalysts, *etc.* A catalyst may markedly increase the velocity of the process, especially in the presence of accelerators and promoters.

Hydrogen peroxide (H₂O₂) undergoes decomposition according to the equation:



The rate is greatly accelerated by various catalysts, *e.g.* manganese dioxide, ferric chloride, or a mixture of ferric chloride and cupric chloride, latter compound acting as a catalytic promoter. The course of the reaction may be followed by titrating aliquots of the mixture with standard solution of potassium permanganate at various intervals of time.

2. Solution Preparation**2.1. 0.02 M KMnO₄ in 250 mL**

$$w = \frac{\text{Molarity} \times \text{Mol. Wt.} \times \text{Volume}}{1000}$$

$$w = \frac{0.02 \times 158.03 \times 250}{1000} = 0.79 \text{ g}$$

Dissolve 0.79 g of KMnO₄ in 250 mL volumetric flask with distilled water and

make upto the mark with distilled water.

2.2. 0.05 M FeCl₃ + 0.04 M HCl in 100 mL

$$w = \frac{0.05 \times 162.2 \times 100}{1000}$$

$$= 0.811 \text{ g FeCl}_3$$

and

$$N_1V_1 = N_2V_2$$

$$11.6 V_1 = 0.04 \times 100$$

$$V_1 = 0.34 \text{ mL HCl}$$

Take 0.8110 g of FeCl₃ and 0.34 mL of concentrated HCl in 100 mL volumetric flask and dissolve in distilled water and make upto the mark with distilled water.

2.3. 1 % (v/v) (0.3 w/v %) H₂O₂ in 250 mL

Dissolve 2.5 mL concentrated H₂O₂ [30 % w/v (or 100 volume)] in 250 mL volumetric flask in distilled water and make upto the mark with distilled water.

2.4. Dilute H₂SO₄ (Appro. 5 N)

Dilute approximate 69.44 mL of concentrated H₂SO₄ to 500 mL with distilled water.

3. Procedure

Place the various solutions in the bath at room temperature for 30 minutes.

Exp. (A) Introduce 100 mL of H₂O₂ 1% solution into a flask, add 5 mL of distilled water and place the flask in the bath. Withdraw immediately 5-mL of the solution, run it into dilute H₂SO₄ and titrate it with standard permanganate solution. Withdraw 5-mL portions after 5, 10, 15, 20, 25, 30, 35, 40, 45, 50, 55 minutes and repeat the above titrations in order to find the amounts of the peroxide left at these intervals of time.

Exp. (B) Carry out a similar experiment using a mixture of 100 mL of 1 % H₂O₂ and 5 mL of the FeCl₃ solution. Withdraw 5-mL portions after 0, 5, 10, 15, 20, 25, 30, 35, 40, 45, 50, 55 minutes and titrate as before.

4. Observation Table

Exp. (A)

Reaction Mixture	
100 mL, 1 % H ₂ O ₂ Solution	5 mL H ₂ O

Aliquot : 5 mL Reaction Mixture

Quenching Solution : Half test tube dilute H₂SO₄

S. No.	Time HH:MM	Time Interval (t) (min)	Burette Reading		$B - A$ (a - x) (mL)	$\log(a - x)$	$\frac{a}{(a - x)}$	$\log \frac{a}{(a - x)}$	$k_A = \frac{2.303}{t} \log \frac{a}{(a - x)}$ (min ⁻¹)
			Initial A (mL)	Final B (mL)					
1	15:05	0	0	10.5	10.5	1.021	1.000	0.000	—
2	15:10	5	0	10.3	10.3	1.013	1.019	0.008	3.847×10^{-3}
3	15:15	10	0	10.3	10.3	1.013	1.019	0.008	1.923×10^{-3}
4	15:20	15	0	10.2	10.2	1.009	1.029	0.013	1.933×10^{-3}
5	15:25	20	0	10.2	10.2	1.009	1.029	0.013	1.450×10^{-3}
6	15:30	25	0	10.1	10.1	1.004	1.040	0.017	1.554×10^{-3}
7	15:35	30	0	9.9	9.9	0.996	1.061	0.026	1.962×10^{-3}
8	15:40	35	0	9.9	9.9	0.996	1.061	0.026	1.681×10^{-3}
9	15:45	40	0	9.8	9.8	0.991	1.071	0.030	1.725×10^{-3}
10	15:50	45	0	9.8	9.8	0.991	1.071	0.030	1.533×10^{-3}
11	15:55	50	0	9.7	9.7	0.987	1.082	0.034	1.585×10^{-3}
							Average k_A		1.919×10^{-3}

Exp. (B)

Reaction Mixture	
100 mL, 1 % H ₂ O ₂ Solution	5 mL, 0.05 M FeCl ₃ Solution

Aliquot : 5 mL Reaction Mixture

Quenching Solution : Half test tube dilute H₂SO₄

S. No.	Time HH:MM	Time Interval (t) (min)	Burette Reading		$B - A$ ($a - x$) (mL)	$\log(a - x)$	$\frac{a}{(a - x)}$	$\log \frac{a}{(a - x)}$	$k_B = \frac{2.303}{t} \log \frac{a}{(a - x)}$ (min ⁻¹)
			Initial A (mL)	Final B (mL)					
1	15:05	0	0	9.5	9.5	0.978	1.000	0.000	—
2	15:10	5	0	7.6	7.6	0.881	1.250	0.097	4.464×10^{-2}
3	15:15	10	0	6.3	6.3	0.799	1.508	0.178	4.108×10^{-2}
4	15:20	15	0	5.2	5.2	0.716	1.827	0.262	4.018×10^{-2}
5	15:25	20	0	4.2	4.2	0.623	2.262	0.354	4.082×10^{-2}
6	15:30	25	0	3.8	3.8	0.580	2.500	0.398	3.666×10^{-2}
7	15:35	30	0	3.2	3.2	0.505	2.969	0.473	3.628×10^{-2}
8	15:40	35	0	2.6	2.6	0.415	3.654	0.563	3.703×10^{-2}
9	15:45	40	0	2.2	2.2	0.342	4.318	0.635	3.658×10^{-2}
10	15:50	45	0	1.9	1.9	0.279	5.000	0.699	3.577×10^{-2}
11	15:55	50	0	1.5	1.5	0.176	6.333	0.802	3.692×10^{-2}
								Average k_B	3.860×10^{-2}

5. Calculation

The values of $(a - x)$, *i.e.* the titration figures, are proportional to the concentrations of the peroxide remaining at time t . Plot $\log(a - x)$ versus t and draw the best straight lines through the point obtained for Exp. (A) and (B) (Fig. 1).

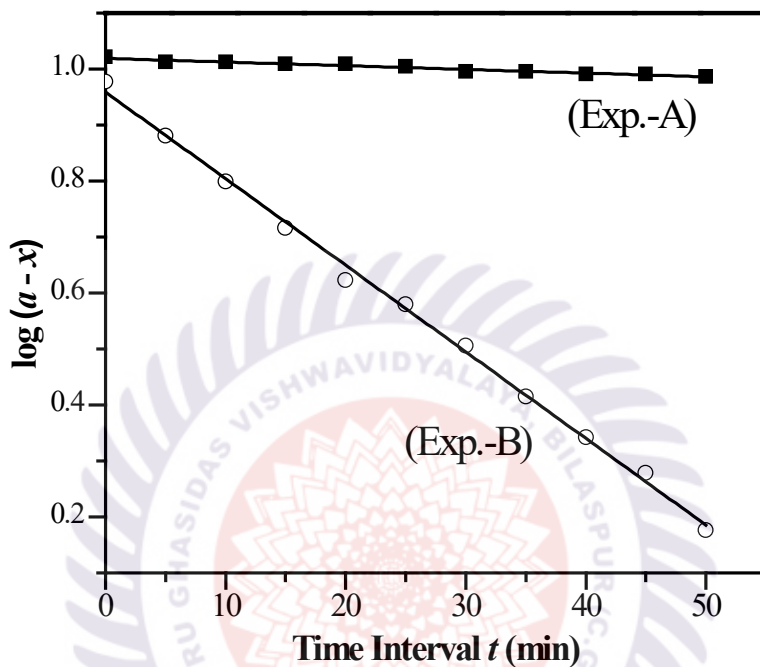


Fig. 1. Plot of $\log(a - x)$ versus time interval t .

Graph shows that the decomposition of H_2O_2 is of the first order under the experimental conditions, *i.e.* it obeys the kinetic equation

$$k = \frac{2.303}{t} \log \frac{a}{(a-x)}$$

where a is the initial concentration of the reactant, and $(a - x)$ is that at time t . Find the slope (S) of the lines in each case and thus calculate the velocity constants.

$$\text{Exp. (A): } S_A = \frac{y_2 - y_1}{x_2 - x_1} = \frac{1.00 - 1.02}{50 - 20} = -0.000666 = -6.7 \times 10^{-4}$$

$$k_A = -S_A \times 2.303$$

$$= 6.7 \times 10^{-4} \times 2.303$$

$$= 1.543 \times 10^{-3} \text{ min}^{-1}$$

$$\text{Exp (B): } S_B = \frac{y_2 - y_1}{x_2 - x_1} = \frac{0.41 - 0.57}{35 - 25} = -0.016$$

$$k_B = -S_B \times 2.303 = 0.016 \times 2.303$$

$$= 0.0368 \text{ min}^{-1}$$

$$= 3.680 \times 10^{-2} \text{ min}^{-1}$$

6. Results and Discussion

The velocity constant, k , of the decomposition of H_2O_2 is found under the experimental conditions as following:

By the graphical method:

$$\text{Rate constant of the Exp. A} = k_A = 1.543 \times 10^{-3} \text{ min}^{-1}$$

$$\text{Rate constant of the Exp. B} = k_B = 3.680 \times 10^{-2} \text{ min}^{-1}$$

By the integration method:

$$\text{Rate constant of the Exp. A} = k_A = 1.919 \times 10^{-3} \text{ min}^{-1}$$

$$\text{Rate constant of the Exp. B} = k_B = 3.860 \times 10^{-2} \text{ min}^{-1}$$

The first-order decomposition of H_2O_2 reaction is probably proceed via an unstable intermediate compound, as in the case of a catalyst consisting of a mixture of a halogen and a halide ion.

7. Precautions

- (i) KMnO_4 should be added drop by drop for titration.
- (ii) H_2O_2 should be carefully diluted.
- (iii) Burette reading should be noted carefully in lower meniscus.

8. Further Reading

- (i) Read up details of the first order kinetic.

- (ii) *Advanced Physico-Chemical Experiments*, J Rose, Sir Isaac Pitman & Sons Ltd, London.
- (iii) *Physical Chemistry*, P. W. Atkins and J. D. Paula, Eds: 7th, Oxford University Press.
- (iv) *Principle of Physical Chemistry*, B. R. Puri, L. R. Sharma, and Pathania, Vishal Publishing Co. Jalandhar.

