**C 141 (Expt. No.______)**
**NAME : ______________________________** **BATCH : _______**
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# STUDY OF THE KINETICS OF THE IODIDE–HYDROGEN PEROXIDE CLOCK REACTION

## AIM

To study the kinetics of the iodide – hydrogen peroxide clock reaction or The Iodine Clock Reaction.

## THEORY

**Iodine clock reactions** are based on the **oxidation of iodide to molecular iodine or the tri-iodide ion.**
In the presence of starch, iodine forms an intense blue colour if the concentration of iodine is of the order of 0.01 millimoles.
In this experiment, the **appearance of the blue colour is retarded by the addition of a reducing agent (sodium thiosulphate)**, which reacts rapidly with the formed iodine or the tri-iodide ion, **converting it back to colourless iodide, until all of the reducing agent is gone**.
The time required for the appearance of the blue colour is inversely related to the initial rate of the reaction, which in turn is also dependent on the temperature and the concentration of the reagents, including the reducing agent.
**This can be expressed as a rate equation, and the order of the reaction can be determined.**
The reaction to be studied is the **acid buffered oxidation of iodide to molecular iodine by hydrogen peroxide**. The formed iodine is made soluble by the iodide ion, in the form of the tri-iodide ion :

$$3\,I^- + H_2O_2 + 2\,H^+ \rightarrow I_3^- + 2\,H_2O \quad \text{----->} \quad \textbf{(slow step)}$$

**<u>The aim is to study the rate of reaction at various concentrations of iodide ion ($I^-$) .</u>**

**The rate of the reaction is measured by determining the time required for the reaction to consume a small amount of thiosulphate, which is initially present in the solution.**
**Thiosulphate is relatively inert towards hydrogen peroxide, but is very rapidly oxidized by iodine :**

$$I_3^- + 2\,S_2O_3^{2-} \rightarrow 3\,I^- + S_4O_6^{2-} \quad \text{----->} \quad \textbf{(fast step)}$$

**No appreciable amount of tri-iodide is formed until the thiosulphate has been completely consumed. At this point, the starch reacts with the tri-iodide, suddenly turning dark blue (starch $I_5^-$ complex). The H+ ion concentration is buffered and remains appreciably unchanged.**

**This is a <u>clock reaction</u>, since the rate of the reaction can be monitored by the <u>sudden appearance of the blue colour</u>.**

## MATERIALS REQUIRED

Standard Volumetric flasks (100 ml); Measuring cylinders (10ml, 25ml); Conical flasks (250ml, 100m l); Burette (50 ml); 1.0 M $H_2SO_4$ solution; 0.001 M sodium thiosulphate solution; 0.4 M potassium iodide solution; 3.0 % hydrogen peroxide solution; 1.0 % starch solution, Distilled Water.

## PROCEDURE

### PART-I

The following solutions **(SOLUTION A and SOLUTION B)** are to be prepared.

**<u>SOLUTION A :</u>**

3.0% $H_2O_2$ (30 ml) + 1.0 M $H_2SO_4$ (30 ml) + 1% starch solution (5 ml) + $H_2O$ (235 ml).

**<u>SOLUTION B :</u>**

The following 5 compositions of SOLUTION B are prepared given in the table below. (Total volume of each will be 100 ml, each contains 10 ml of 0.001 M sodium thiosulphate.)

Table

| S.No. | 0.001 M $Na_2S_2O_3$ (ml) | 0.4 M KI (ml) | $H_2O$ (D.W.) (ml) | Total Volume (ml) | Conc. Of KI (M) | Conc. Of $H_2O_2$ (M) | Time (sec) | | | Mean time (t) (sec) | (1/t) $sec^{-1}$ |
|---|---|---|---|---|---|---|---|---|---|---|---|
| | | | | | | | $t_1$ | $t_2$ | $t_3$ | | |
| $B_1$ | 10.0 | 10.0 | 80.0 | 100.0 | | | | | | | |
| $B_2$ | 10.0 | 15.0 | 75.0 | 100.0 | | | | | | | |
| $B_3$ | 10.0 | 20.0 | 70.0 | 100.0 | | | | | | | |
| B4 | 10.0 | 25.0 | 65.0 | 100.0 | | | | | | | |
| B5 | 10.0 | 30.0 | 60.0 | 100.0 | | | | | | | |

PART-II

1. Mix 25 ml of Solution A with 20 ml of Solution B1 in a 100 ml conical flask, and start the stop-watch immediately. Gently swirl once and keep it constant on the bench, and note down the time required for the first appearance of the blue colour.
2. Repeat the reaction two more times for the B1 solution.
3. Repeat the procedure with $B_2$, $B_3$, $B_4$ & $B_5$ solutions also, by noting the 3 readings in each case.
4. Calculate the average time 't' (sec) and the reciprocal of the average reaction time "1/t" ($sec^{-1}$) for each set.
5. Calculate the molar concentrations of KI, $Na_2S_2O_3$, and $H_2O_2$ used in each case. Assume the density of the used $H_2O_2$ solution (3.0%) as 1.01 g / ml.

DISCUSSIONS

1. In all the solutions, different Iodide concentrations are present, but there is a fixed amount of thiosulphate ions. Thus, a fixed amount of iodine is consumed before the appearance of the blue color.
2. The reaction times of the different sets tells us how long it takes to liberate a fixed amount of iodine, for varying iodide concentrations.
3. **"Rate"** is directly related to a **constant** divided by the **time it takes to reach the end point.**

**OBSERVATIONS AND CALCULATIONS**

1. Determine the Relative Rate for each reaction mixture, by using the reciprocal of the average reaction time (1/t).
2. **Plot a graph between 1/t against [ $I^-$ ] and the slope gives the value of k'.**
3. The rate equation will be the above chemical reaction will be in the form,

$$\textbf{rate} = k [H_2O_2]^p [I^-]^n [H^+]^q$$

4. But, here $[H_2O_2]^p$ and $[H^+]^q$ are keeping constants in the experiment, and so this simplifies to:

**Rate = k' [ $I^-$ ]n**
**(where k' is a constant other than k)**

**log(Rate) = log(k') + n . log[ $I^-$ ]**

**Rate α 1/t (or) Rate = k''/ t**

**and [$I^-$] α volume of $I^-$ solution**

**Hence, log(Rate) = log(k'') – log(t) ; k'' is a constant.**

**log[( Rate)y / Rate)x] = log( tx / ty)**

**log[( Rate)y / (Rate)x] = n . log([ $I^-$ ]$_y$ / [ $I^-$ ]$_x$)**
**Therefore, n = log( tx / ty) / log([ $I^-$ ]$_y$ / [ $I^-$ ]$_x$)**

5. **Calculate the 'n' value for each solution B( $B_1$,$B_2$,$B_3$,$B_4$ & $B_5$).**
6. Thus, a plot of the **Rate** versus **[ $I^-$ ]** gives an indication of the exponent, **'n'** .
7. The value of 'n' would give the order of the reaction with respect to the iodide ions.
8. To calculate the unit of **'n'**, use the formula: **$(lit)^{n-1}$ $(mole)^{1-n}$ $(sec)^{-1}$.**

**RESULTS**

1. Rate constant of the reaction, **k'** = …………………........ at …....... $^oC$.
2. The order of the reaction, **'n'** is ........……………..........
3. Comment on the nature of the graph(s).