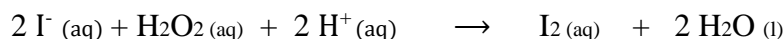


### Answer the Following Exercise:

#### Study of a Slow Reaction

Hydrogen peroxide ( $\text{H}_2\text{O}_2$ ) oxidizes iodide ions ( $\text{I}^-$ ) in acidic medium in a slow reaction which takes place according to the following equation:



The aim of this exercise is to study the kinetic of this reaction.

#### 1. Preparation of a Hydrogen Peroxide ( $\text{H}_2\text{O}_2$ ) Solution ( $\text{S}_1$ )

Available is a hydrogen peroxide solution ( $\text{S}_0$ ) of concentration  $C_0 = 2.7 \text{ mol.L}^{-1}$ . It is required to prepare a hydrogen peroxide solution ( $\text{S}_1$ ) of concentration  $C_1 = 0.1 \text{ mol.L}^{-1}$ .

Choose, from the **document-1**, the most suitable set of glassware to realize this dilution. Justify.

Set A	Set B	Set C
Volumetric flask of 100 mL Volumetric pipet of 10 mL Beaker of 50 mL	Volumetric flask of 100 mL Graduated pipet of 10 mL Beaker of 50 mL	Volumetric flask of 100 mL Graduated cylinder of 10 mL Beaker of 50 mL

**Document-1**

#### 2. Preliminary Study

In a beaker, one mixes:

- A volume  $V_1 = 18 \text{ mL}$  of potassium iodide solution ( $\text{K}^+ \text{I}^-$ ) of concentration  $C_1 = 0.1 \text{ mol.L}^{-1}$
- A volume  $V_2 = 9 \text{ mL}$  of sulfuric acid solution  $\text{H}_2\text{SO}_4$  of concentration  $C_2 = 1 \text{ mol.L}^{-1}$ .

At instant  $t = 0 \text{ s}$ , a volume  $V_3 = 3 \text{ mL}$  of hydrogen peroxide solution  $\text{H}_2\text{O}_2$  of concentration  $C_3 = 0.1 \text{ mol.L}^{-1}$  is added to the beaker.

In this mixture, sulfuric acid is in excess.

- 2.1. Calculate the initial concentrations of iodide ions  $[\text{I}^-]_0$  and hydrogen peroxide  $[\text{H}_2\text{O}_2]_0$  in the reaction mixture.
- 2.2. Show that hydrogen peroxide  $\text{H}_2\text{O}_2$  is the limiting reactant.
- 2.3. Calculate  $[\text{I}_2]$  at the end of experiment.

#### 3. Kinetic Study

By an appropriate method, the concentration of iodine  $[\text{I}_2]$  at different instants is determined. The results are grouped in the table of **document-2**.

t (s)	100	200	300	400	500	600	650
$[\text{I}_2] (10^{-3} \text{ mol.L}^{-1})$	3.85	5.9	7.5	8.6	9.4	9.85	10

**Document-2**

**3.1.** Plot the curve representing the variation of the concentration of iodine as a function of time  $[I_2] = f(t)$  in the interval of time  $[0 - 650 \text{ s}]$ .

Take the scales: In abscissa: 1 cm for 50 s ;

In ordinates: 1 cm for  $1 \times 10^{-3} \text{ mol.L}^{-1}$ .

**3.2.** Deduce, graphically, the variation in the rate of formation of  $I_2$  as a function of time.

**3.3.** Show that the molar concentration of  $I_2$  and that of  $I^-$  ions, at instant  $t$ , are related by the following relation:

$$[I_2]_t = 3 \times 10^{-2} - \frac{[I^-]_t}{2}$$

**3.4.** Verify if  $t = 650 \text{ s}$  represents the end time of reaction.

**3.5.** Determine the average rate of formation of  $I_2$  in the interval of time  $[200 - 500 \text{ s}]$ .

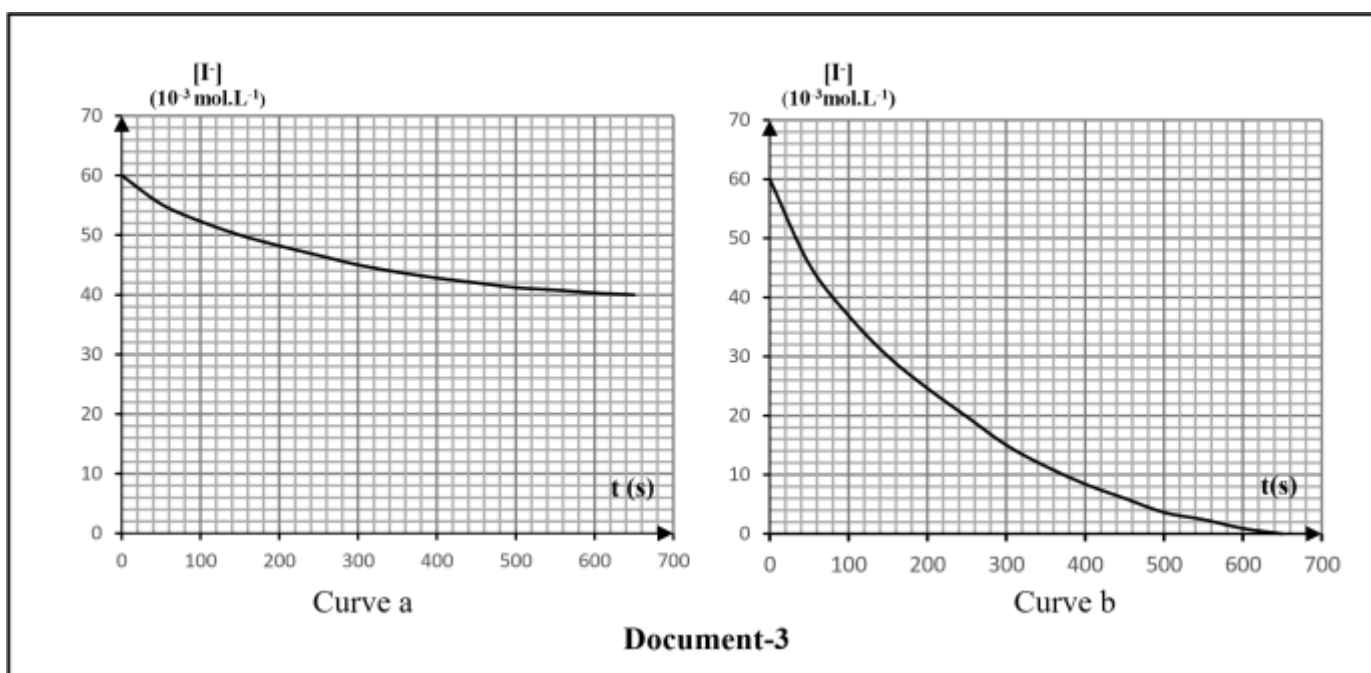
**3.6.** Deduce the average rate of disappearance of  $I^-$  in the same interval.

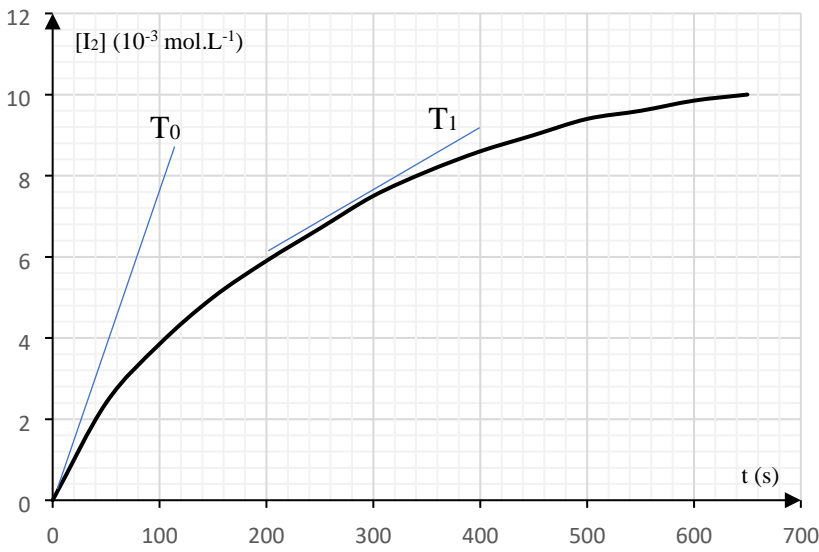
**3.7.** Justify the following statements:

**3.7.1.** The initial rate of formation of  $I_2$  is greater than its rate of formation at  $t = 300 \text{ s}$ .

**3.7.2.** The concentration of iodide ions at  $t = 300 \text{ s}$  is  $[I^-]_{t=300} = 45 \times 10^{-3} \text{ mol.L}^{-1}$ .

**3.8.** From the curves of **document-3**, specify the one that corresponds to the variation of the concentration of iodide ions as a function of time  $[I^-] = g(t)$ .



Exercise 1 (10 points) Study of a Slow Reaction		
Part of Q	Answer	Mark
1	<p>During dilution, the number of moles of solute is conserved</p> $C_0V_0 = C_f V_f$ <p>For <math>V_f = 100 \text{ mL}</math>, <math>V_0 = \frac{C_f \times V_f}{C_0} = \frac{0.1 \times 100}{2.7} = 3.7 \times 10^{-3} \text{ L} = 3.7 \text{ mL}</math>, therefore a graduated pipet must be used. So we choose <b>Set B</b></p>	1
2.1	<p>For <math>\text{I}^-</math> : <math>n_1 = C_1 \times V_1 = 0.1 \times 18 \times 10^{-3} = 1.8 \times 10^{-3} \text{ mol}</math></p> <p>For <math>\text{H}_2\text{O}_2</math> : <math>n_2 = C_2 \times V_2 = 0.1 \times 3 \times 10^{-3} = 0.3 \times 10^{-3} \text{ mol}</math></p> <p>• <math>[\text{I}^-]_0 = \frac{n_1}{V_{\text{total}}} = \frac{1.8 \times 10^{-3}}{30 \times 10^{-3}} = 0.06 \text{ mol.L}^{-1}</math>.</p> <p>• <math>[\text{H}_2\text{O}_2]_0 = \frac{n_2}{V_{\text{total}}} = \frac{0.3 \times 10^{-3}}{30 \times 10^{-3}} = 0.01 \text{ mol.L}^{-1}</math>.</p>	1
2.2	<p><math>R(\text{I}^-) = \frac{n_1}{2} = \frac{1.8 \times 10^{-3}}{2} = 0.9 \times 10^{-3}</math></p> <p><math>R(\text{H}_2\text{O}_2) = \frac{n_2}{1} = 0.3 \times 10^{-3}</math></p> <p><math>R(\text{I}^-) &gt; R(\text{H}_2\text{O}_2)</math> hence <math>\text{H}_2\text{O}_2</math> is the limiting reactant.</p>	0.75
2.3	<p>Calculate the concentration of iodine at the end of the reaction: According to the stoichiometric ratios:</p> $\frac{n_0(\text{H}_2\text{O}_2)}{1} = \frac{n(\text{I}_2)_\infty}{1}, \text{ so } n(\text{I}_2)_\infty = 0.3 \times 10^{-3} \text{ mol}$ $[\text{I}_2]_\infty = \frac{n(\text{I}_2)_\infty}{V_{\text{total}}} = \frac{0.3 \times 10^{-3}}{30 \times 10^{-3}} = 0.01 \text{ mol.L}^{-1}$	0.75
3.1		1
3.2	<p>The instantaneous rate of formation of <math>\text{I}_2</math> is equal to the slope of the tangent drawn to the curve at time of abscissa <math>t</math></p> <p>Slope of tangent <math>T_0 &gt;</math> slope of tangent <math>T_1</math></p> <p>As time increases, the slope of tangent decreases thus the rate of reaction decreases.</p>	1
3.3	<p>According to the stoichiometric ratios: <math>\frac{n(\text{I}^-)_{\text{reacted}}}{2} = \frac{n(\text{I}_2)_{\text{formed}}}{1}</math></p> <p><math>n(\text{I}^-)_{\text{remaining}} = n_0(\text{I}^-) - n(\text{I}^-)_{\text{reacted}}</math>     <math>n(\text{I}^-)_{\text{remaining}} = n_0(\text{I}^-) - 2n(\text{I}_2)_{\text{formed}}</math> divide by total volume</p>	1

	$[I^-]_t = [I^-]_o - 2[I_2]_{\text{formed}} \quad 2[I_2]_t = [I^-]_o - [I^-]_t$ $\text{Then } [I_2]_t = \frac{[I^-]_o}{2} - \frac{[I^-]_t}{2} \quad [I_2]_t = \frac{0.06}{2} - \frac{[I^-]_t}{2}$ $\text{then } [I_2]_t = 3 \times 10^{-2} - \frac{[I^-]_t}{2}$	
<b>3.4</b>	$[I_2]_{\infty} = 0.01 \text{ mol.L}^{-1} = 10 \times 10^{-3} = [I_2] \text{ at } t = 650\text{s}$ So, $t = 650\text{s}$ represents the end of the reaction	<b>0.5</b>
<b>3.5</b>	Average rate of formation of $I_2$ is equal to the positive slope of secant drawn to the curve between $t=200\text{s}$ and $t=500\text{s}$ $r_{t200 \rightarrow 500} = \frac{\Delta[I_2]}{\Delta t} = \frac{(9.4 - 5.9) \times 10^{-3}}{500 - 200} = 1.16 \times 10^{-5} \text{ mol.L}^{-1} \text{ s}^{-1}$	<b>1</b>
<b>3.6.</b>	According to stoichiometry: $\frac{r_{I_2}}{1} = \frac{r_{I^-}}{2} \text{ at same time } t \quad r_{I^-} = 2 \times 1.16 \times 10^{-5} = 2.32 \times 10^{-5} \text{ mol.L}^{-1} \text{ s}^{-1}$	<b>0.5</b>
<b>3.7.1</b>	The rate of formation of iodine decreases with time due to the decrease in the concentration of reactants.	<b>0.5</b>
<b>3.7.2</b>	Refer to the table: $[I_2]_{t=300} = 7.5 \times 10^{-3} \text{ mol.L}^{-1}$ . Substitute in the relation: $[I_2]_t = 3 \times 10^{-2} - \frac{[I^-]_t}{2} \quad [I^-]_{t=300} = 2 \times 3 \times 10^{-2} - 2 \times 7.5 \times 10^{-3} = 45 \times 10^{-3} \text{ mol.L}^{-1}$	<b>0.5</b>
<b>3.8</b>	<b>Curve a</b> Since $[I^-]_o = 60 \times 10^{-3} \text{ mol.L}^{-1}$ $[I^-] \text{ at } t = 650 \text{ s}$ is not zero because $I^-$ is an excess reactant.	<b>0.5</b>