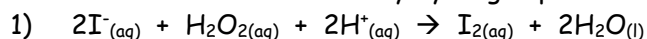


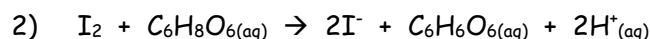
Study of Reaction Rates: Clock Reaction Lab

What is the effect of concentration on the rate of a reaction? In this experiment, a special type of chemical reaction- called a clock reaction- will be used to determine quantitatively the influence of concentration on rate.

The first reaction used is the oxidation of iodide ions by hydrogen peroxide in aqueous solutions:



This reaction is relatively slow. However, the reaction of iodine with vitamin C to produce iodide ions is relatively fast.



In a solution containing iodide, hydrogen peroxide and vitamin C, we find that iodine will not accumulate until all the vitamin C is used up, because iodine from reaction (1) will immediately be used up in reaction (2). The addition of starch to the mixture will indicate when this point occurs, because starch forms a deep blue complex with iodine. When all the vitamin C is gone, therefore, the mixture will suddenly turn blue. A given set of concentrations will give reproducible reaction rate, hence the name clock reaction.

In this experiment, the amount of vitamin C will be the same in each reaction mixture. The time required for the blue colour to appear will signal how long it takes to produce the same amount of iodine (see second equation). This time will therefore represent the rate of the reaction (1). By then varying the concentrations of iodide and hydrogen peroxide, we can thus determine how this affects the rate of the reaction.

Objectives:

- 🌀 To determine the time required to produce a given amount of iodine in a clock reaction
- 🌀 To determine how concentration of reactants affects the rate of a clock reaction
- 🌀 To present the results graphically
- 🌀 To determine the rate law for reaction (1)

Apparatus and Materials:

beakers	magnetic stir bars
stirring plate	stop watch
volumetric pipets	
0.20 M potassium iodide, KI - freshly made	
3% hydrogen peroxide (0.88 M)	
Starch solution	
Vitamin C solution (ascorbic acid) - 2x500 mg tablets dissolved in 60 mL water - freshly made	

Procedure:

A: Studying the Effect of [KI] on Rate of Reaction

1. Fill 10 beakers according to the chart below:

Beaker #	Volume of KI (mL)	Volume of H ₂ O (mL)	Volume Vitamin C (ml)		Beaker #	Volume of H ₂ O ₂ (mL)	Volume of starch (mL)
1A	50	0	5.0		1B	10	2.5
2A	40	10	5.0		2B	10	2.5
3A	30	20	5.0		3B	10	2.5
4A	20	30	5.0		4B	10	2.5
5A	10	40	5.0		5B	10	2.5

2. Place beaker 1A on the stir plate. Insert a magnetic stir bar into the beaker. Pour the contents of beaker 1B into 1A and immediately begin timing. Stop the timer with the first appearance of colour. Record the time in the observation chart.
3. Repeat with the remaining beakers.

B: Studying the Effect of [H₂O₂] on Rate of Reaction

4. Repeat the above procedure filling the beakers as outlined below:

Beaker #	Volume of KI (mL)	Volume Vitamin C (ml)		Beaker #	Volume of H ₂ O ₂ (mL)	Volume of water (mL)	Volume of starch (mL)
1A	50	5.0		1B	10	0	2.5
2A	50	5.0		2B	8	2	2.5
3A	50	5.0		3B	6	4	2.5
4A	50	5.0		4B	4	6	2.5
5A	50	5.0		5B	2	8	2.5

Observations from Part A:

Beaker #	Time (s)
1	
2	
3	
4	
5	

Observations from Part B:

Beaker #	Time (s)
1	
2	
3	
4	
5	

Calculations:

1. For each mixture, calculate the concentrations of iodide and hydrogen peroxide just before the reaction begins. The starting concentrations of potassium iodide and hydrogen peroxide are given at the beginning of the lab. By considering the volume of KI and H_2O_2 that you placed in each reaction beaker and the fact that the total volume placed in each beaker was 67.5 mL and using the dilution formula: $C_1V_1 = C_2V_2$ you can calculate the initial concentrations of I^- and H_2O_2 . Record in the charts below.

Show 1 sample calculation:

A: Studying the Effect of $[\text{I}^-]$ on Rate of Reaction

$[\text{I}^-]$ (mol/L)	$[\text{H}_2\text{O}_2]$ (mol/L)	Time (s)

B: Studying the Effect of $[\text{H}_2\text{O}_2]$ on Rate of Reaction

$[\text{I}^-]$ (mol/L)	$[\text{H}_2\text{O}_2]$ (mol/L)	Time (s)

2. Plot two graphs: Iodide concentration vs reaction time (horizontal axis)
Hydrogen peroxide concentration vs reaction time
3. For each graph, draw tangents at three different concentrations, the second being twice the first and the third three times the first. i.e 2M, 4M and 6M.
4. Calculate the instantaneous rate at each tangent. Show all work on graph.

5. Complete the charts below:

$[I^-]$ (mol/L)	Rate (Mol/L·s)

$[H_2O_2]$ (mol/L)	Rate (Mol/L·s)

6. Showing your work on the charts above as practiced in class, determine the order of the reaction with respect to iodide and to hydrogen peroxide. The numbers that you get won't be as perfect as we used in class so you will need to use a little discretion.

7. What is the rate law for the reaction?

Questions:

1. In procedure A, what did we do to keep the $[H_2O_2]$ the same in each trial? Why was it necessary to keep $[H_2O_2]$ the same?
2. Another way to determine the order of a reaction can be seen on page 378 in your text book. Study each of the graphs on page 378 and then describe in your own words how this method works.
3. Plot the reciprocal of the reaction time (a measure of the average rate) - on the y-axis- against the concentration of iodide.
Repeat with the concentration of hydrogen peroxide.
What does each graph tell you about the order of each reactant?
Does the order of each reactant agree with what you determined earlier?

$[I^-]$ (mol/L)	$1/t$ (s^{-1})

$[H_2O_2]$ (mol/L)	$1/t$ (s^{-1})