THE IODINE CLOCK REACTION

A2 4.1/A

Aim

Hydrogen peroxide reacts with iodide ions in acid solution to form iodine.

$$H_2O_2(aq) + 2\,H^+(aq) + 2\,I^-(aq) \rightarrow I_2(aq) + 2\,H_2O(l)$$

The rate equation for this reaction has the form: $\text{rate} = k\,[H_2O_2]^m\,[I^-]^n\,[H^+]^p$

By varying the concentration of the iodide ion, you can determine the order of this reaction with respect to the concentration of the iodide ions (i.e. find *n*).

The liberated iodine is reacted with thiosulphate(VI) ions, which have been added to the reaction mixture.

$$I_2 + 2\,S_2O_3^{2-} \rightarrow 2\,I^- + S_4O_6^{2-}$$

When all of the thiosulphate ions have been used up, free iodine remains in solution and this is detected by the formation of a blue-black colour with starch indicator. The appearance of the blue-black colour represents the same extent of reaction in each case, and so the initial rate is proportional to 1/time.

Background

This section goes beyond the specification but explains the principle behind the experiment.

The rate equation for this reaction will be of the form: $\text{rate} = k\,[H_2O_2]^m\,[I^-]^n\,[H^+]^p$

In your experiment you will be keeping $[H_2O_2]$ and $[H^+]$ constant, and so this simplifies to:

$$\text{rate} = K\,[I^-]^n \quad \text{(where K is a different constant to k)}$$

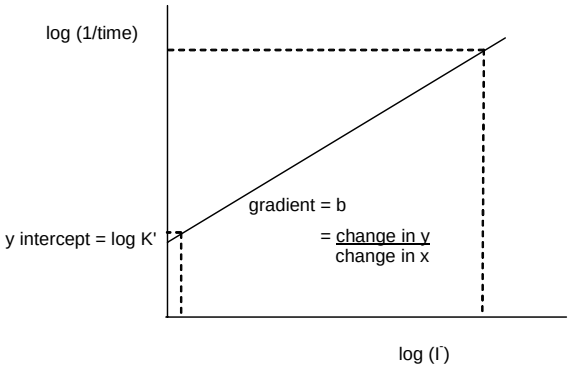
Taking logs: $\log \text{rate} = n \log [I^-] + \log K$

The rate is proportional to 1/time, and $[I^-]$ is proportional to the volume of I^- solution.

$$\log(1/\text{time}) = n \log(\text{volume } I^-) + \log K' \quad (K' \text{ is a different constant to K})$$

This equation represents a straight line graph (y = mx + c).

Therefore a plot of log (1/time) versus log (volumeI^-) should give a straight line graph of gradient *n*, the order with respect to $[I^-]$.



Safety


sulphuric acid


iodine

Method

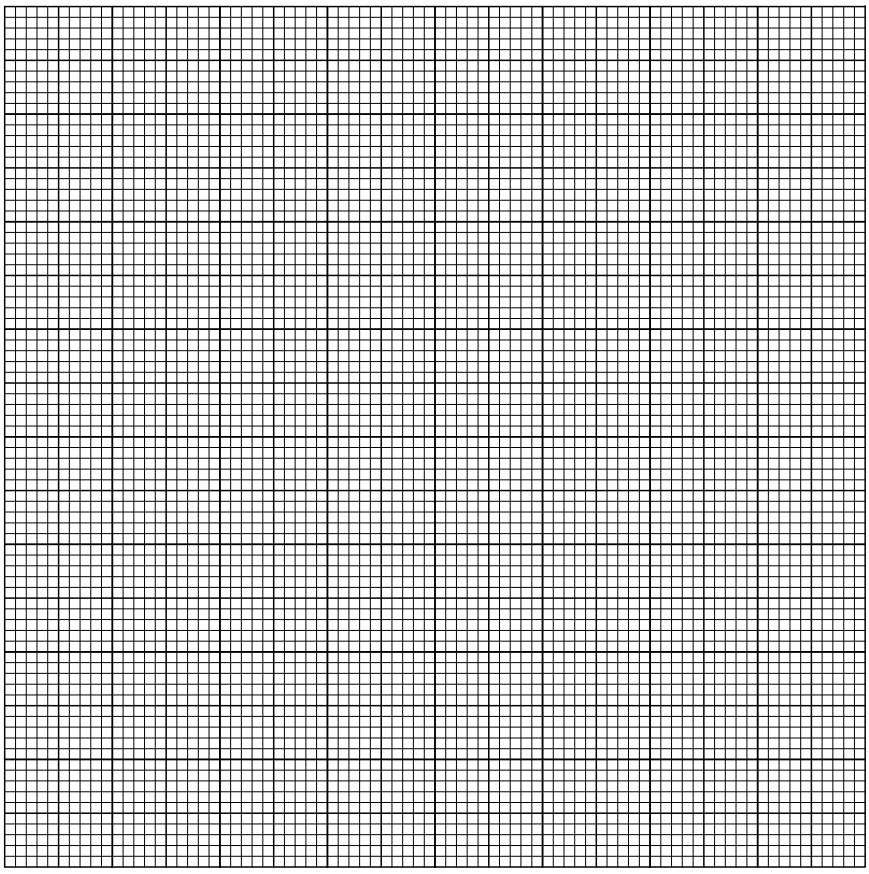
In each experiment,

1) Use pipettes and burettes as shown to mix the following volumes of solutions in a small clean conical flask.
2) Use the burette to place 10 cm^3 of 0.10 mol dm^{-3} hydrogen peroxide in a test tube.
3) Pour the hydrogen peroxide from the test tube into the flask, start timing immediately and swirl the contents to mix thoroughly.
4) Record the time for the blue colour of the starch-iodine complex to appear in the table.

Experiment	sulphuric acid (0.25 mol dm^{-3})	starch	water	potassium iodide (0.10 mol dm^{-3})	sodium thiosulphate (0.002 mol dm^{-3})	time (s)
	25 cm^3 pipette	teat pipette	burette	burette	pipette	
A	25	1	20	5	10	
B	25	1	15	10	10	
C	25	1	10	15	10	
D	25	1	5	20	10	
E	25	1	0	25	10	

Analysis

1) Use your results and the method outlined above to determine the order of the reaction with respect to iodide ions. There is space for a table of data and then the graph grid.



Gradient = ……..

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Order with respect to I^- = ……

2) Assuming that the maximum errors for the apparatus in this experiment are as shown below, estimate the percentage error in measuring out the volume of **each separate reagent** in experiment A, and **then** the overall percentage error. The starch was present as an indicator and so its exact volume was not vital.

burette $\pm$ 0.1 cm^3 25 cm^3 pipette $\pm$ 0.1 cm^3 10 cm^3 pipette $\pm$ 0.1 cm^3

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Evaluation

3) Look at your graph. Is your line of best fit good enough for you to deduce an order with confidence?

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4) Are there any anomalous results? Suggest one possible reason for an anomalous result, even if all your results are within tolerance.

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5) The reaction is first order with respect to the concentration of the iodide ion. Calculate the percentage difference between your value for the order and the correct value.

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6) Are your results accurate? Explain your answer. ……

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7) Beside errors in measuring volumes, identify two other sources of error in this experiment. Suggest one improvement to minimise each source of error.

a) ……..

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b) ……..

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