# Using the iodine clock method to find the order of a reaction assignment

Using the Iodine clock method to find the order of a reaction Introduction When peroxodisulfate (VI) ions and iodide ions react together in solution they form sulfate (VI) ions and iodide. This reaction is shown below: $S_2O_8^{2-}aq+ 2I^-aq$ $SO_4^{2-}aq+ I_2(aq)$ The reactants and the sulfate (VI) ions are colourless however the Iodine is a yellow/brown colour. This allows you to measure the progress of the reaction through the colour change when the iodine is produced. In order to determine the order of the reaction we need to measure the initial rate of the reaction by timing how long the reaction takes to produce a small, fixed amount of iodine.

This can be measured clearly and effectively by using the iodine clock method. When you add thiosulfate (VI) ions to the starting reaction mixture the thiosulfate (VI) ions turn iodine back into iodide ions. The equation for this reaction is shown below: $2S_2O_3^{2-}aq+ I_2aq$ $S_4O_6^{2-}aq+ 2I^-(aq)$ The products of this reaction are both colourless so no Iodine colour will appear until all the thiosulfate (VI) ions have been used up. The appearance of the iodine can be made much clearer by adding some starch to the reaction mixture as iodine forms an intense black/blue colour in its presence.

This means that the mixture will suddenly experience a colour change from colourless to intense blue/black. By measuring the time taken for this colour change to occur you know how long it took to use up all the thiosulfate (VI) ions and so how long it took to produce the iodine that reacted with it. This is known as the iodine clock method and can be used to find the order of a reaction and determine the effect of Iodide ion concentration on rate of reaction. The procedure First of all I measured out 5 different volumes of 1.

00 mol dm?? aqueous potassium iodide solution and put these into 5 different boiling tubes.

This was the only reactant that differed in concentration throughout the investigation. All other reactant concentrations were kept constant and water was added to the mixtures in order to maintain a constant volume of solution throughout. Next I added the aqueous 0. 0100 mol dm?? sodium thiosulfate (VI) solution to each boiling tube. I added 2. 00 cm? of this each time. I then added 1. 00cm? of starch solution to the mixture in each boiling tube. In order to keep the volume of the mixture in each boiling tube constant I added enough water to each mixture to bring the volume up to 10cm?.

I measured the temperature of the solution in each boiling tube with a thermometer to ensure that each reaction occurred at the same temperature as difference in temperature could have an effect on the rate of reaction. After these mixtures were complete I could then add the potassium peroxodisulfate (VI) solution to each mixture containing varying KI concentrations. I added 2. 00cm? of potassium peroxodisulfate (VI) solution to the first mixture and started timing immediately, using a stopwatch, until the colour change was observed. I used the thermometer to continuously stir the solution.

I then repeated this process with the remaining for mixtures and recorded the time taken each time. The table below shows the quantities used in each mixture: Mixture| Volume of KI (aq)/cm? | Volume of water/cm? | Volume of Na? S? O? (aq)/cm? | Volume of starch solution/cm? | Volume of K? S? O?

(aq)/cm? | 1| 5. 00| 0. 00| 2. 00| 1. 00| 2. 00| 2| 4. 00| 1. 00| 2. 00| 1. 00| 2. 00| 3| 3. 00| 2. 00| 2. 00| 1. 00| 2. 00| 4| 2. 00| 2. 00| 2. 00| 1. 00| 2. 00| 5| 1. 00| 4. 00| 2. 00| 1. 00| 2. 00| To measure out all the solutions I used a 5cm? burette ensuring the burette was rinsed thoroughly between different reactants to avoid contamination.

I measured the solution so that the meniscus was touching the graduation at eye level, ensuring the tip was full, and kept this constant throughout. I also ensured that I touched the side of the boiling tube with the end of the burette each time to ensure the last drop of solution was fully drained out of the burette. These were all efforts to ensure reliable results. Results and Calculations The table below shows the results that I collected: Mixture| Concentration of I? (aq)/mol dm?? | Clock time/s| Rate/moldm?? s?? | Temp/?? C| 1| 0. 5| 26. 00| 0. 038| 27| 2| 0. 4| 39. 34| 0. 025| 27| 3| 0. 3| 59. 59| 0. 017| 27| 4| 0. 2| 104. 47| 9. 7X10?? | 27| 5| 0. 1| 293. 38| 3. 41X10?? | 27| The graph attached shows the progress curve for this experiment. Calculations In order to calculate the concentration of Iodide ions in the solution I used the equation: [I-] = Volume of KI cm 3 x [KI] mol dm3 Total Volume of Mixture An example of how I calculated the concentration of mixture 5 is shown below: Mixture 5 [I? ] = 1. 0cm3 X 1. 0mol dm-310cm3 = 0. 5 moldm?? In order to calculate the concentration of [S? O??? ] ions in the solution I used the formula: Concentration = Volume of K2S2O8 X K2S2O8Total volume of mixture I did this for mixture 5: Mixture 5 [S? O??? ] = 2. 0cm3X 0. 04molddm-310cm3 = 8X10?? mol dm?? By comparing these two concentrations for mixture 5 I was able to identify that the iodide ions were in excess. I used mixture 5 as it contains the lowest concentration of KI

which means it can be easily compared to the $K_2S_2O_8$ solution which was kept constant. Next I calculated the total amount of iodine produced in the reaction. This can be done by looking at the mole to mole ratio in the reaction between peroxodisulfate (VI) ions and Iodide ions, shown below: $S_2O_8^{2-}(aq) + 2I^-(aq) \rightarrow SO_4^{2-}(aq) + I_2(aq)$ The equation shows that there is a 1: 1 mole ratio between the $S_2O_8^{2-}$ ions and the $I_2$ produced.

This suggests that Moles of $S_2O_8^{2-}$ = Moles of $I_2$. As previously calculated, the concentration of $[S_2O_8^{2-}]$ ions in the solution is $8 \times 10^{-3}$ mol dm$^{-3}$ suggesting that there is the potential for $8 \times 10^{-3}$ mol dm$^{-3}$ of Iodine to be produced. Next I calculated the number of moles of thiosulfate ions added to each mixture. To do this I used the formula: Concentration= Volume of $Na_2S_2O_3$ X $Na_2S_2O_3$Total volume of mixture = 2. 0cm3X 0. 01mol dm-310cm3 = $2 \times 10^{-5}$ I then calculated the amount of iodine used up by the thiosulfate ions. I did this by using the mole to mole ratio from the equation of the reaction between thiosulphate ions and iodine.

This is shown below: $2S_2O_3^{2-}(aq) + I_2(aq) \rightarrow S_4O_6^{2-}(aq) + 2I^-(aq)$ The equation shows that 2 moles of $S_2O_3^{2-}$ react with 1 mole of $I_2$ – in other words the reaction is a 2: 1 mole ratio. We already know that the number of moles of thiosulphate ions used in the reaction is $2 \times 10^{-5}$ mol dm$^{-3}$. Therefore, according to the ratio the number of moles of iodine used is half of this: $1 \times 10^{-5}$ mol dm$^{-3}$. I was then able to identify the percentage of the reaction studied. To do this I used the formula: % Reaction studied = Moles of Iodine usedMoles of Iodine made X 100 = 1. 0X10-38X10-3X 100 = 12. % Calculating Initial Rate In order to calculate the initial rate of the reaction I used the formula: Rate= Number of Moles of Iodine usedTime take for colour

to appear or Rate= 1Time The results of these calculations can be seen in the table on page 2. Rate Equation: After plotting the progress curve for this investigation, see attached graph, I could determine that the reaction is first order in respect to iodide ions as the line of the graph was straight. This can be written as: Rate= k[I? ] [S? O??? ] This means that the overall order of the reaction is 2.

In order to calculate the rate constant (K) I calculated the gradient of the line of best fit on the graph. This followed the basic formula that: K= Gradient= Change in yChange in x = rate[I-] From my graph I calculated that at 0. 5 mol dm?? K= 0. 032s-10. 5 mol dm-3 = 0. 064 K is the rate constant and is constant for a constant temperature. The rate constant will increase with an increase in temperature. Conclusion My investigation shows that as the concentration of Iodide ions increases the rate of reaction also increases. I have also calculated that the reaction is first order in respect to iodide ions and is second order overall.