Investigating the rate of reaction between peroxydisulphate (vi) ions and iodide ...

Aims The aims of this investigation are: 1. To find the rate equation of the reaction of hydrogen peroxide and iodide ions. This will be achieved by using an iodine clock method and colorimetric analysis. 2. Draw a graph of rate against concentration for each reactant (Hydrogen peroxide, potassium iodide and H+ ions). 3. Finding the order for each reactant 4. Finding the rate-determining step. 5. Proposing a mechanism for the reaction. 6. Using Arrhenius' equation to find the activation enthalpy. Background The basic reaction for this can be illustrated with the following equation: $3I^-(aq) + H_2O_2(aq) + 2H^+(aq) \rightarrow I_3^-(aq) + 2H_2O(aq)$ (1) The half equations for this reaction can be written as follows: $3I^- \leftrightarrows I_3^- + 2e^-$ $H_2O_2 + 2H^+ + 2e^- \leftrightarrows 2H_2O$ This reaction demonstrates that iodide ions are oxidised by hydrogen peroxide to tri-iodide ions. This is stage one of a sequence of reactions, which continues below: $I_3^-(aq) + 2S_2O_3^{2-}(aq) \neg \leftrightarrows 3I^-(aq) + S_4O_6^-(aq)$ (2) This shows that the tri-iodide ions are reduced back to iodide ions by the thiosulphate ions. Thus, the iodine that is formed in reaction (1) is immediately transformed into iodide ion and we do not see the blue-black colour of the starch-iodide complex until all of the thiosulphate ion has reacted with $I_2(aq)$ and is exhausted. $I_3^-(aq) + starch \leftrightarrows Starch-I_5^-$ complex $+ I^-(aq)$ Once the thiosulphate ion has been exhausted, the tri-iodide ion can react with the starch, forming the Starch-I5- complex, giving the blue-black colour. When this occurs, we will then know the amount of hydrogen peroxide that has reacted and the time it took to react. These equations will, thus, enable the slow step (rate-determining step) to be determined, which is another aim of this experiment. Though details of the starch and iodine reaction are not yet fully known, it is thought that iodine fits inside the coils

of amylose. The transfer of charge between the iodine and the starch and the spacing between the energy levels in the complex formed corresponds to the absorption spectrum, and so, the complementary colour, a blue-black solution, is observed. Factors that affect the rates of reaction , There are many factors that affect the rate of a reaction. These include surface area, concentration difference, presence of a catalyst, pressure and temperature. Affect of surface area on the rate of reaction Surface area can also affect the rate of reaction. A reaction will happen more quicker if the solid is finely divided into a powder, rather than a lump of the same mass. This is because a reaction can only occur if the particles taking part in the reaction collide. A larger surface area provides a higher likelihood of collisions (and thus, a reaction) to take place. One example is called the " Bread and Butter Theory." If you take a loaf of bread and cut it into slices it, you have more surfaces to spread butter onto. Taking a more practical example, In the above example, reactant 1 can get to the outer atoms of reactant 2, but not the central atoms. This has a small surface area. However, if the surface area is increased, Reactant 1 can get to all the atoms of reactant 2. Increasing the number of collisions per second increases the rate of reaction. Affect of concentration on the rate of reaction For many reactions, including this one, increasing the concentration of the reactants increases the rate of the reaction. This is because, for a reaction to occur, a collision must take place first. Increasing the concentration of the reactants will increase the frequency of the collisions between two reactants, as there are a higher number of reactants to collide with. From a probabilistic point of view, if there are a higher number of reactants (i. e. a higher concentration), the

chance of a collision, and therefore, a reaction to take place, increases. For example, if we have the following situation: supposing fixed positions and an equal probability of being hit, the probability of a green particle hitting a red particle is 1/3. If we increase the number of red particles to 2, the probability now of a green particle hitting a red particle is ½, which is thus, an increase by 1/6. Although the temperature is being kept constant, however, the kinetic theory is applicable. This is because the molecules involved in the reaction have a range of energy levels. When colliding molecules have sufficient energy, a reaction takes place. If they do not, then a reaction cannot take place. This is because the temperature that is being measured is only the average temperature (and thus, kinetic energy, because T α Ek) of the substance. It is impossible for the kinetic energy of every atom in the substance to be the same, and so, the temperature is an average. The reason for this is that each molecule has a certain amount of kinetic energy and once it collides (perfectly elastically) with another molecule, it transfers its kinetic energy to the molecule it collided with, giving it a higher kinetic energy than the initial molecule. In most cases, when you increase the concentration, the rate of reaction also increases. In certain multi-step reactions, however, the reaction happens in a series of small steps. Suppose the reaction happens as so: The speed at which A splits into X and Y dictates the rate of the reaction. This is also known as the rate-determining step. If you increase the concentration of A, the chances of the first step happening also increase, due to the increase in the number of molecules of A. Increasing the concentration of B undoubtedly speeds up the second step, but makes little difference to the overall rate. Affect of the presence of a

catalyst on the rate of reaction A catalyst is a substance that speeds up a reaction by providing an alternative pathway with a lower activation enthalpy, and is chemically unchanged at the end of a reaction. Reactions can only take place if the two reactants collide with enough energy to initiate the reaction (i. e., to begin breaking the bonds). Majority of the molecules do not have enough energy, and simply bounce apart after collisions. One way of speeding up a reaction is to provide an alternative pathway for the reaction to occur with a lower activation energy. In other words, the activation energy on the Maxwell-Boltzmann Distribution graph should look like this: A catalyst does this, and this can be shown on an enthalpy profile diagram: A catalyst works in two ways. One of them is adsorption, and this is where the molecules are attached to the surface of the catalyst due to the weak interactions (typically Van Der Waal's forces) between the surface and the reactants. Initially, the bonds in the reactants weaken and break. Bonds are then formed between the reactants, forming the products, and they then diffuse away from the surface of the catalyst. Another way is the formation of the intermediate compounds, and many catalysts, including all enzymes, work by forming intermediates. The reactants involved in the reaction combine with the catalyst making an intermediate compound, but this is very unstable. When this intermediate breaks down, it releases the new compound and the original catalyst. Affect of the pressure on the rate of reaction Increasing the pressure on a reaction involving gases increases the rate of reaction. This does not happen with reactions involving solids or liquids. Increasing the pressure works in the same way as increasing the concentration; if the pressure of a given mass is increased, it is just

squashed into a smaller volume. Having the same number of particles in a smaller volume works in the same way as increasing the concentration. The ideal gas equation illustrates this (As liquids act similarly to gases, the ideal gas equation can give a fair demonstration of how liquids would act): $pV = nRT$ Where p = pressure, V = volume n = number of moles R = molar gas constant T = Temperature (in Kelvin) This can be re-arranged to read: $p = n/v \times RT$ n/v is the number of moles divided by the volume, which is the concentration. RT is constant at a constant temperature. Because this is just a constant, it can be shown that $p = k(n/v)$ or $p \propto (n/v)$ (the pressure is proportional to the concentration). Thus, if you double one, the other will also be doubled. Whether you are considering a reaction where collisions between two different particles or two of the same particles occur, the same law applies: for any reaction to occur, collisions must happen first. This is true when both particles or one of the two sets of particles are in the gaseous state. If the pressure is higher, the chance for collisions to occur also increases. If the reaction involves a particle splitting, the reacting particles must have enough energy to react. Supposing that one in a hundred particles have enough energy to react - if you had two hundred particles, two of them would react, and so forth, so if you double the pressure, the rate of reaction also doubles. The main variable to be tested in this experiment will be the concentration of hydrogen peroxide and the concentration of iodide ions. These will be varied in two separate experiments, thus enabling a fair test to be attained. If both variables are altered at once in an experiment, it would be very difficult to say which variable has had more effect on the rate of reaction. Therefore, in the first

investigation the concentration of hydrogen peroxide will be varied, keeping all other variables constant and in the second investigation the concentration of iodide ions will be altered, keeping all the other variables constant. The results of this investigation will enable me to draw a rate graph for the two investigations. This will later enable me to combine these two rates to form an overall rate equation. The iodine clock reaction is a chemical reaction in which two colourless solutions are mixed and react together to form a red/brown colour. However, initially, the iodine will be of a small concentration, and will appear very light in colour, and therefore, the production of iodine will be very hard to detect. This is, thus, enhanced by the addition of starch, which instantaneously turns dark blue/black with the formation of iodine ions, giving a more accurate time for the production of iodine ions. A colorimeter is an instrument that measures the concentration of a substance by comparing its colour with the standard (i. e. distilled water). In this experiment, the complementary colour to the orange of iodine is blue-green, so a blue-green filter (470 nm) will be used in method two on the colorimeter. This reaction demonstrates that reaction rates depend on the concentrations of the reagents involved in the overall reaction. The time required to reach this point depends on the rates of the first two reactions, and consequently on the concentrations of all the reactants. Anything that accelerates the first reaction (e. g., iron catalysis or temperature) will shorten the time. Thus, increasing the concentration of iodide, hydrogen peroxide, or acid (it neutralises the hydroxide ion) will accelerate the reaction. On the other hand, increasing the thiosulphate concentration will have the opposite effect; it will take longer for the blue colour to appear. The

rate of reaction is a measure of how fast the reaction occurs. The graph that can be drawn from the results is time against volume (concentration) of the variable solution. A rate of reaction can then be obtained. Average rates are not very good comparisons, because the reaction may finish before the designated time interval. The fairest way to measure the rate is at the start, as a