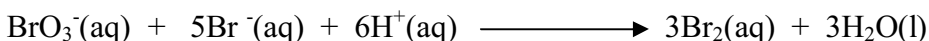


Concentration and Temperature Affect the Rate of Reaction

Aim

I aim to investigate how the concentration of potassium bromide affect the rate of reaction when added to a solution made up of potassium bromate(V), sulphuric acid and phenol.

The general equation for the reaction between bromide and bromate ions in acidic aqueous solution is: ⁽¹⁾



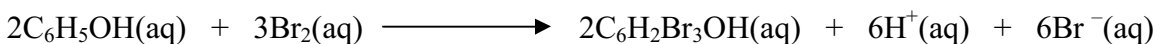
I am going to alter the concentrations of potassium bromide, potassium bromate and sulphuric acid to find the orders of reaction with respect to each reactant. This will allow me to prove that the rate equation for this reaction is: ^(1, pg 230)

$$\text{Rate} = k[\text{BrO}_3^-][\text{Br}^-][\text{H}^+]^2$$

I will be working out a value for the rate constant, k, and will also be investigating the affect temperature has on rate of reaction, using my results and the Arrhenius equation to work out the activation enthalpy.

Theory:

The colour change in this reaction is from an orange, brought about by the methyl orange indicator, to a colourless solution. As bromine is produced in the reaction it, at first, bonds to the phenol. However, when all binding sites on phenol have been used up, the presence of excess bromine will turn the solution colourless. This shows that I am measuring the time taken for all binding sites on phenol to be used up, which explains the change from orange to colourless in the solution. This can be shown in the following equation:



The term 'rate of reaction' refers to how fast reactants are converted into products in a given reaction. A way to measure rate of reaction directly has not yet been discovered, and therefore we must measure the change in amount of product/reactant in a certain time. This can be shown by the following equation: ^(1, pg 228)

$$\text{Rate} = \frac{\text{Change in Property}}{\text{Time Taken}}$$

However, orders of reaction must also be taken into account when writing a rate equation. Where A and B are reactants, m and n are orders of reaction (with respect to A and B) and k is the rate constant:

$$\text{Rate} = k[A]^m[B]^n$$

Reactions can be either zero, first or second order with respect to each molecule used in the reaction. A reaction with order zero will be a horizontal line on a graph of concentration against rate of reaction. First order will be a straight diagonal line and second order will show a curve. This is shown in Figure 4 below: ⁽²⁾

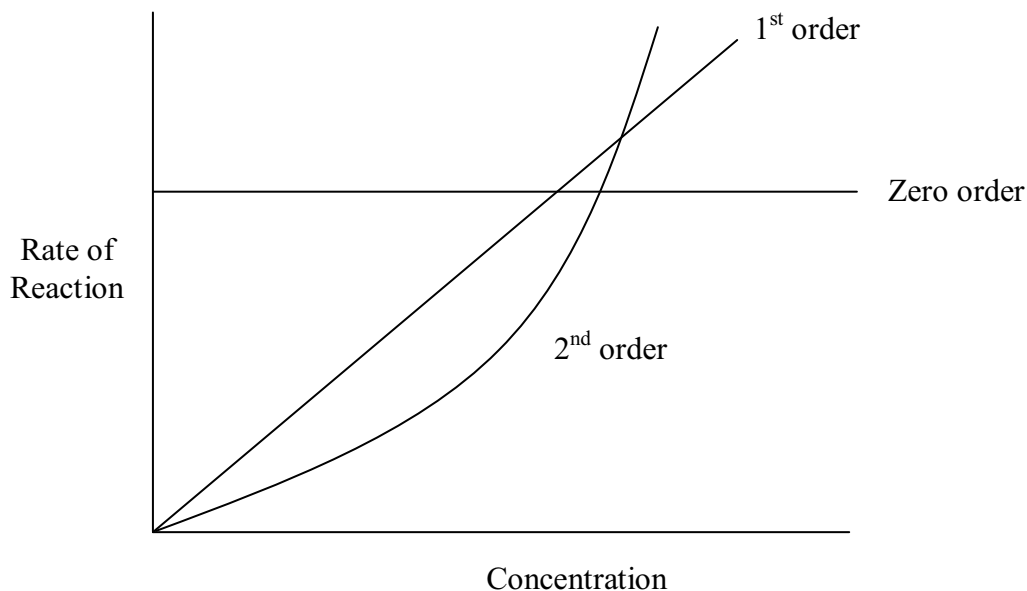


Figure 4 – model graph showing orders of reaction

For a reaction with order zero:

$$\text{Rate} = k[A]^0$$

This means, due to the fact that anything to the power of zero equals one, the equation for a zero order reaction is:

$$\text{Rate} = k$$

This shows that, for order zero reactions, the rate is independent of the concentration of 'A'. It has no effect on the rate of reaction. If this is the case, the reactant in question is excluded from the rate equation, as it has no bearing on it.

In some reactions changing the concentration of a particular reactant has no effect on the rate. When this is the case it means there are two steps involved in the reaction - this is known as the reaction mechanism. Where there are two or more steps it is the slowest step that determines the overall rate. This is called the rate determining step. I will suggest a mechanism later in my investigation for the reaction, based on the rate equation.

The rate equation for the reaction I am investigating is shown in my 'Aim' section. This shows that the reaction is expected to be first order with respect to both bromate ions and bromide ions, and second order with respect to hydrogen ions.

The rate constant in a rate equation, shown as k , is not affected by a change in concentration but is affected by a change in temperature. This means that the temperature of an equation must always be known when working out the rate of reaction.

Several different things contribute to a change in rate of reaction which concerned in my experiment:

- Concentration
- Temperature
- Collision Frequency
- Activation Enthalpy

The Collision Theory can account for changes in both concentration and temperature affecting rate of reaction. This theory suggests that reactions occur when the particles of reactants collide, provided that they collide with a certain minimum kinetic energy.

Concentration of Reactants:

If the concentration of the reactant molecules is increased there will be more molecules. This means that the likelihood of collisions occurring is increased, so a reaction will happen sooner (as shown in figure 3). This increased the 'rate of reaction' as a reaction occurs in a shorter space of time. The diagrams below show this effect: ^(1, pg 221)

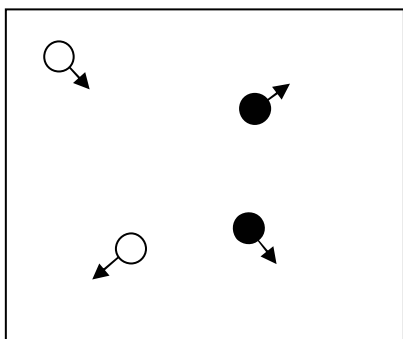


Figure 2 – lower concentration of reactants reduces the chances of a collision.

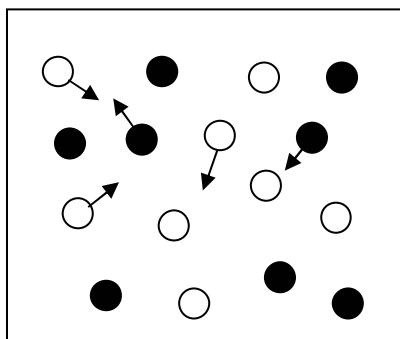


Figure 3 – higher concentration of reactants means that a collision is much more likely.

Temperature:

When molecules collide, they must do so with enough energy for the reaction to take place. If the temperature of the reactants is increased, the energy of the molecules will increase. This also means that the molecules will move at a faster speed, meaning that collisions are likely to occur sooner. For many reactions happening at room temperature (roughly 20-25°C), it is known that the rate of reaction doubles each time the temperature is increased by 10°C.^(1, pg 224) However, this is not true for all equations and even equations where this is true the value may be slightly different, for example 9 or 11 rather than 10°C. Also, the temperature change needed to double the rate of reaction will vary gradually as temperature is increased.

Increasing the Collision Frequency:

The frequency of a collision between two molecules in a reaction is proportional to the square root of the temperature in Kelvin.⁽³⁾ For example, if the temperature is increased from 293K to 303K, a change from 20 to 30°C, the factor increase can be calculated as follows:

$$\sqrt{\frac{300}{310}} = 1.016$$

This is an increase of 1.6% for a temperature change of +10°C. As I have already mentioned, the rate of reaction will have approximately doubled with this rise in temperature – this is around a 100% increase. This shows that the effect of increasing collision frequency on the rate of reaction is almost negligible.

Activation Enthalpy: (1, pg 224/225)

The Activation Enthalpy is described as the minimum amount of kinetic energy needed before a collision will occur between two molecules. This energy is needed for bonds in the reactants to be broken and new bonds in the products to be formed.

An important factor in the rate of reaction is how much energy the particles involved have. This energy is increased when the temperature of the reaction is increased. The Maxwell-Boltzmann distribution is a good way of showing this:

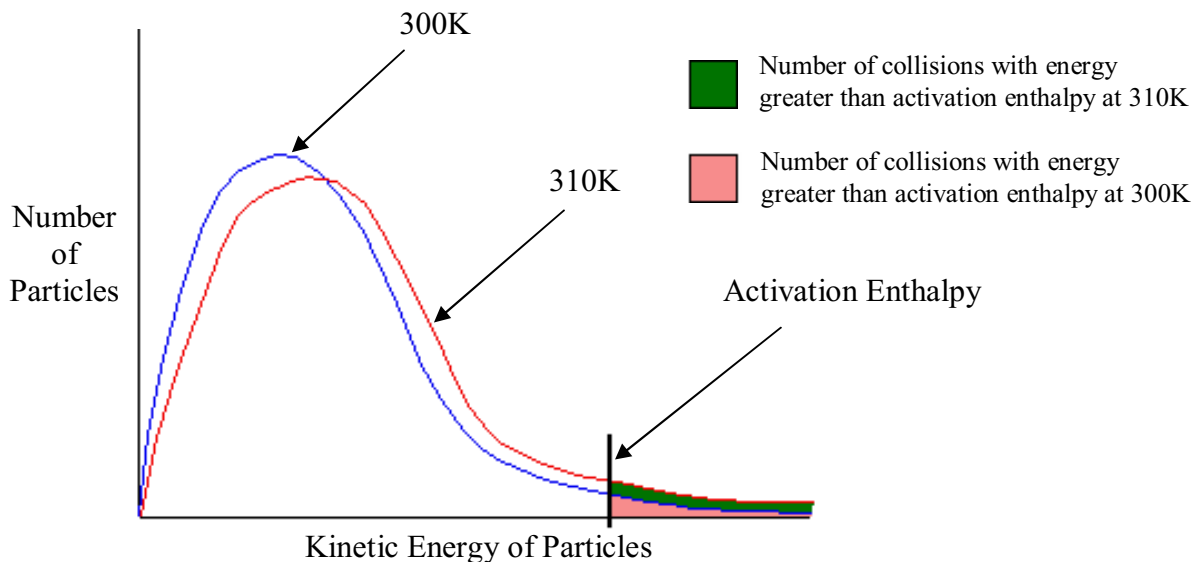


Figure 4 – The Maxwell-Boltzmann Distribution

Figure 3 above shows that as temperature increases, more molecules move at higher speeds and have a higher kinetic energy. Figure 4 shows the change when the temperature is increased by 10K. It shows that there is still a wide range of energies but now there are more particles with a higher kinetic energy. At the higher temperature of 310 Kelvin there are many more molecules which have reached the activation enthalpy level than at 300K. The amount of molecules reaching the activation enthalpy is about twice as many at 310K than at 300K, so this means that the reaction will go around twice as fast at 310K in comparison to the reaction taking place at 300K.

Put simply, “reactions go faster at higher temperatures because a larger proportion of the colliding molecules have the minimum activation enthalpy needed to react.”

The Arrhenius Equation ⁽⁴⁾

Where:

k = rate constant

A = Arrhenius constant (also known as frequency factor)

e = mathematical value similar to pi (has a value of 2.71828)

E_a = the activation enthalpy

R = the gas constant (has a value of 8.31 J K⁻¹ mol⁻¹)

T = temperature (in Kelvin)

$$k = Ae^{-E_a/RT}$$

This equation can be rearranged by taking the log (to the base e) form of the equation. Log to the base e is denoted by ln:

$$\ln k = \ln A - \frac{E_a}{RT}$$

This can be further rearranged to give:

$$\ln k = \frac{-E_a}{R} \times \frac{1}{T} + \ln A$$

If a graph of ln k against 1/T is drawn, it should form a straight line of best fit with gradient -E_a/R.

This gradient can be measured and the activation enthalpy can be worked out.

Making Molar Solutions ⁽⁵⁾

Molar (M) solutions are based on the number of moles of chemical in 1 litre of solution. A mole consists of 6.02x10²³ molecules or atoms. Relative Formula Mass (RFM) is the weight of one mole of a chemical. This can be determined using a periodic table by adding the atomic mass of each atom in the chemical formula. I have shown below how to work out the RFM of a compound, using 'X' and 'Y' as elements, with 'n' and 'm' as their respective molecular masses:

X₂Y₃:

$$X_2 = 2 \times n = 2n$$

$$Y_3 = 3 \times m = 3m$$

$$= 2n + 3m \text{ g/mole}$$

Therefore, a 1M solution of X₂Y₃ consists of 2n+3m grams of X₂Y₃ dissolved in enough water to make one litre of solution.

Once the molecular weight of a chemical is known, the weight (in grams) of chemical to dissolve in water for a molar solution less than 1M is can be calculated using the following formula:

$$\text{grams} = \frac{(\text{desired molarity of solution}) \times (\text{RFM of chemical}) \times (\text{desired amount of solution/ml})}{1000\text{ml}}$$

For example, to make 100 ml of 0.1M X_2Y_3 solution, you would work out how many grams of X_2Y_3 to use as is shown below:

$$\text{grams of } \text{X}_2\text{Y}_3 = \frac{[0.1 \times (2n3m) \times 100]}{1000}$$

This would give you the amount of grams of X_2Y_3 required, with the rest of the solution being made up with distilled water.

Pre-test

Firstly I have decided to use burettes rather than measuring cylinders for measuring out my solutions. This is due to the fact that they are more accurate apparatus. The error associated with a 50cm^3 burette is $\pm 0.1\text{cm}^3$. This is far more accurate than the value for a 25cm^3 measuring cylinder, which is $\pm 0.5\text{cm}^3$. When using a burette the measurement should be read from the bottom of the meniscus. It should also be read at eye level to avoid a parallax error.

Each measurement in my actual experiment will be carried out 5 times to confirm my result.

Repetitions increase reliability by proving that each reading is not a 'one off' and can be achieved the majority of the time. An average can then be taken from the 5 repetitions. This average time can be used to work out reaction rate.

Volumes of 5cm^3 have been used to keep the percentage errors to a minimum, although at the same time making sure that I have enough solution that I can see any colour change easily. This means that my percentage error for the burettes is as follows:

$$\text{Percentage error} = \frac{\text{error}}{\text{measured value}} \times 100$$

So for measuring, for example, 5cm^3 from a burette the error would be:

$$\text{Percentage error} = \frac{0.1}{5} \times 100$$

$$= 2\%$$

I need to decide upon the range of concentrations of potassium bromide I will use in my experiment. To do this I will test concentrations of 0.001, 0.002, 0.003, 0.004, 0.005, 0.006, 0.008 and 0.01 mol/dm⁻³ of potassium bromide. I will then see how long each test took and decide how many of these trials I can do in the time I have available.

Secondly I need to find the best way of stirring the mixture whilst the reaction occurs. Preferably, I need a way of stirring which will reduce the time taken for the solution to go colourless. This would mean that I can fit in more repeats and more concentrations of potassium bromide can be tested. However, I also need a method which does not stir the mixture so vigorously that the colour change cannot be seen easily.

Pre-test Results:

Concentration of Potassium Bromide (mol/dm ⁻³)	Time Take For the Mixture to Turn Colourless (seconds)
0.001	293.0
0.002	131.5
0.003	84.0
0.004	68.0
0.005	60.5
0.006	56.5
0.008	44.5
0.01	38.5

Although all of these concentrations are shown to have worked, I do not have enough time available to conduct results including each concentration. If I am going to repeat each concentration several times for more accurate results I need to investigate a smaller range of concentrations. I have chosen to exclude the two lowest concentrations (0.001 and 0.002 mol/dm⁻³) when conducting my actual experiment. This is because they took the longest time for the mixture to turn colourless. Whilst conducting this pre-test I used a magnetic stirrer. However, I was not sure whether this was the most appropriate method to use. To solve this I conducted another pre-test, which is shown in the table below:

Concentration of Potassium Bromide (mol/dm ⁻³)	Time Taken When Stirred by Hand (seconds)	Time Taken When Stirred by Magnetic Stirrer (seconds)	Time Taken When Left Without Stirring (seconds)
0.003	76.0	67.5	60.5
0.004	51.5	52.0	53.5
0.005	41.5	45.5	39.5
0.006	36.0	40.5	44.5
0.008	33.0	29.5	33.5
0.01	31.5	24.0	25.0

From my results a reliable trend can be seen from the results obtained when stirring by hand and when using the magnetic stirrer. However, when the mixture is left without stirring anomalies arise, as is shown in the table. There is a reliable trend until the

concentration of 0.005mol/dm^{-3} , where the time taken for the mixture to turn colourless decreases rather than increasing. To ensure that I could see the colour change easily, I placed the conical flask on a white tile. This meant that the colour change would be more obvious. I have decided to use the magnetic stirrer as it allows me to concentrate more on when the colour change occurs, rather than stirring the mixture. When I was stirring the mixture by hand, I found that I had a tendency to lose concentration and lose track of when the colour change occurred. Having a magnetic stirrer meant that all I needed to do was focus on the mixture and note the timing of the colour change. I believe this will help me in obtaining accurate results.

Temperature of Reactants (°C)	Time Take For the Mixture to Turn Colourless (seconds)
10	101.5
20	23.7
30	10.2
40	5.1
50	1.9
60	0.8

I also conducted a pre-test to trial my method for investigating how temperature change affects the reaction rate. I came across a few problems that need to be adjusted in my final experiment. I originally decided to just monitor the temperature of the water bath, but soon realised that this would most likely not share the same value as the solutions in the boiling tubes. To overcome this problem I placed a thermometer in each boiling tube, one in the tube containing potassium bromide and one in the solution of the other reactants. Once both had reached the correct temperature I was able to mix the two solutions, knowing that both were at the required temperature. I also discovered that when investigating temperatures above 40°C the reaction became too fast to record accurate timings. I would have lowered the concentration of potassium bromide, although I wanted to keep this constant to my previous experiments investigation how concentration affects rate of reaction. Instead, I chose the following range of temperatures: 10, 15, 20, 25, 30, 35 and 40°C . This range also makes it easier for me to keep the temperature of the reactants stable, as they do not fluctuate to far above or below room temperature.

For the concentrations of the other reactants used in this reaction I decided to alter them by the same amount and in the same range as I have chosen to alter the concentration of potassium bromide. This ensures that the experiment is as fair as possible, as all reactants are being varied by the same amount. It could be harder to work out a rate equation if I vary the concentrations of the reactants by different amounts to each other.

Method for Investigating Concentration Change

Apparatus:

Apparatus	Reason For Use	Size/Concentration
Burettes	To store solutions before adding to mixture. This ensures that the solutions are kept under the same conditions as each other, and that accurate amounts of solution can be measured out.	50ml
Potassium bromate(V) solution	Used as part of the reaction.	0.01M
Sulphuric acid solution	Used as part of the reaction.	1M
Phenol solution	Used as part of the reaction.	0.0001M
Methyl orange indicator	Used to show colour change in the reaction as phenol binding sites are used up by bromine.	n/a
Potassium bromide	Used as part of the reaction.	0.01M
Distilled water	Used to dilute solutions to chosen concentrations.	n/a
Beakers	Used for reaction to take place in.	250ml
White tile	Allows easy observation of colour change against a white background.	n/a
Magnetic stirrer	Allows me to concentrate on observing the colour change alone, rather than having to conduct stirring and observe colour change at the same time.	n/a
Stopwatch	Used to measure the time taken for the solution to turn colourless.	n/a
Volumetric flask	Used to make up solutions accurately.	250ml

Method

1. Take 5 burettes and label them as “potassium bromate(V)”, “sulphuric acid”, “phenol”, “potassium bromide” and “distilled water”. Wash them out with water and then with the corresponding solution. Fill each burette with its corresponding solution.
2. Add 5cm³ of potassium bromate(V) solution, 5cm³ of sulphuric acid solution and 5cm³ of phenol solution to a beaker. Add 4 drops of methyl orange indicator to the beaker. These amounts should be kept constant throughout the experiment to ensure that the test is as fair as possible. The same solution should be used in each trial to keep the test fair, as different solutions may have very slight differences in concentration.
3. Add 5cm³ of potassium bromide solution of the required concentration to another beaker. For most trials the potassium bromide will need mixing with distilled water to reach the desired concentration.
4. Place the first beaker onto a white tile and pour into it the potassium bromide solution. Start the stopwatch and continuously stir the mixture until it goes completely colourless. Record the time taken for this to happen.
5. Wash the beakers out with distilled water and repeat the experiment using concentrations of 0.003, 0.004, 0.005, 0.006, 0.008 and 0.01mol/dm⁻³ of potassium bromide. Repeat each concentration 5 times. All other factors such as temperature should be kept constant and the same apparatus should be used each time. Sulphuric acid concentration should be kept at 0.1mol/dm⁻³ when not being investigated, so that it is not so weak that the reaction takes place too slowly to be measured. All other solutions should be kept at 0.01mol/dm⁻³.
6. Once all concentrations of potassium bromide solution has been investigated and repeated, conduct experiments with respect to potassium bromate(V) and sulphuric acid. To do this, simply alter the concentration of the chosen reactant as was done when using potassium bromide. Concentration can be changed simply by taking a sample of the original solution and adding distilled water. Where M_1 is the molarity of the solution before dilution, M_2 is molarity after dilution, V_1 is the volume of the solution before dilution and V_2 is volume after dilution:

$$V_1 = \frac{(M_2 \times V_2)}{M_1}$$

For example, to make a 250ml solution with concentration 0.008M of potassium bromide from a solution of 0.01M, you would do the following:

$$\begin{aligned} V_1 &= \frac{(0.008 \times 250\text{ml})}{0.01} \\ &= 200\text{ml} \end{aligned}$$

This means that you would need to use 200ml of the original 0.01M solution with 50ml of distilled water to make 250ml of 0.008M solution.

Remember to keep all other solutions at a constant concentration throughout.
Investigate the same range of concentrations as was used for potassium bromide,
to ensure that a rate equation can easily be found.

Making Required Solutions

0.01M potassium bromate(V):

KBrO_3 :

$$\text{K} = 1 \times 39.1 = 39.1$$

$$\text{Br} = 1 \times 79.9 = 79.9$$

$$\text{O}_3 = 3 \times 16.0 = 48.0$$

$$\text{RFM} = 167$$

$$\text{grams needed} = \frac{(0.01 \times 167 \times 250)}{1000}$$

= 0.4175g dissolved in enough distilled water to make 250ml of solution.

1M Sulphuric acid:

H_2SO_4 :

$$\text{H}_2 = 2 \times 1.0 = 2$$

$$\text{S} = 1 \times 32.1 = 32.1$$

$$\text{O}_4 = 4 \times 16.0 = 64.0$$

$$= 98.1$$

$$\text{grams needed} = \frac{(1 \times 98.1 \times 250)}{1000}$$

= 24.525g dissolved in enough distilled water to make 250ml of solution.

0.0001M phenol:

$\text{C}_6\text{H}_5\text{OH}$:

$$\text{C} = 6 \times 12.0 = 72.0$$

$$\text{H} = 6 \times 1.0 = 6.0$$

$$\text{O} = 1 \times 16.0 = 16.0$$

$$= 94.0$$

$$\text{grams needed} = \frac{(0.0001 \times 94 \times 250)}{1000}$$

= 0.00235g dissolved in enough distilled water to make 250ml of solution.

0.01M potassium bromide:

KBr:

$$K = 1 \times 39.1 = 39.1$$

$$\begin{aligned} Br &= 1 \times 79.9 = 79.9 \\ &= 119.0 \end{aligned}$$

$$\text{grams needed} = \frac{(0.01 \times 119 \times 250)}{1000}$$

= 0.2975g dissolved in enough distilled water to make 250ml of solution.

Risk Assessment ⁽⁶⁾

The following risk assessments are based on concentrated solutions of each substance. The concentrations I will be using are much more diluted than the solutions mentioned below to ensure that my experiment is safer. Although some of the acute hazards and symptoms associated with the concentrations I am using will not be as severe as those shown below, the risks are still very serious. For this reason I will still take appropriate precautions and will handle each substance with care.

Potassium bromate(V):

Type of Hazard	Acute Hazards/Symptoms	Prevention of Hazard	Solving Hazard
Fire	Not combustible but enhances combustion of other substances. Gives off irritating or toxic fumes (or gases) in a fire.	No contact with combustibles and reducing agents.	Water in large amounts.
Explosion	Risk of fire and explosion on contact with combustible substances and reducing agents.	n/a	In case of fire: keep drums etc. cool by spraying with water.
Inhalation	Cough. Sore throat.	Breathing protection.	Fresh air, rest. Refer for medical attention.
Contact with skin	Redness.	Protective gloves.	First rinse with plenty of water, then remove contaminated clothes and rinse again. Refer for medical attention.
Contact with eyes	Redness. Pain.	Safety goggles, or eye protection in combination with breathing protection if powder.	First rinse with plenty of water for several minutes (remove contact lenses if easily possible), then take to a doctor.
Ingestion	Abdominal pain. Diarrhoea. Nausea. Vomiting.	Do not eat, drink, or smoke during work. Wash hands before eating.	Rinse mouth. Give a mixture of activated charcoal in water to drink. Induce vomiting if person is conscious. Refer for medical attention.
Spillage	n/a	n/a	Sweep spilled substance into sealable containers; if appropriate, moisten first to prevent dusting. Carefully collect remainder then remove to safe place. Do not absorb in sawdust or other combustible absorbents.
Storage	n/a	Separated from combustible and reducing substances, powdered metals and incompatible materials. See Chemical Dangers.	n/a

Sulphuric acid:

Type of Hazard	Acute Hazards/Symptoms	Prevention of Hazard	Solving Hazard
Fire	Not combustible. Many reactions may cause fire or explosion. Gives off irritating or toxic gases in a fire.	No contact with flammable substances. No contact with combustibles.	No water. In case of fire in the surroundings: powder, foam, carbon dioxide.
Explosion	Risk of fire and explosion on contact with bases, combustible substances, oxidants.	n/a	In case of fire: keep drums etc. cool by spraying with water but no direct contact with water.
Inhalation	Corrosive. Burning sensation. Sore throat. Cough. Struggling to breathe. Shortness of breath.	Ventilation or breathing protection.	Fresh air, rest. Half-upright position. Artificial respiration may be needed. Refer for medical attention.
Contact with skin	Corrosive. Redness. Pain. Blisters. Serious skin burns.	Protective gloves. Protective clothing.	Remove contaminated clothes. Rinse skin with plenty of water or shower. Refer for medical attention.
Contact with eyes	Corrosive. Redness. Pain. Severe deep burns.	Face shield or eye protection in combination with breathing protection.	First rinse with plenty of water for several minutes (remove contact lenses if easily possible), then take to a doctor.
Ingestion	Corrosive. Abdominal pain. Burning sensation. Shock or collapse.	Do not eat, drink, or smoke during work.	Rinse mouth. Do not induce vomiting. Refer for medical attention.
Spillage	n/a	n/a	Do not absorb in saw-dust or other combustible absorbents. Personal protection: complete protective clothing including self-contained breathing apparatus. Do not let this chemical enter the environment.
Storage	n/a	Separated from combustible and reducing substances, strong oxidants, strong bases, food and feedstuffs, incompatible materials. May be stored in stainless steel containers. Store in an area having corrosion resistant concrete floor.	n/a

Phenol:

Type of Hazard	Acute Hazards/Symptoms	Prevention of Hazard	Solving Hazard
Fire	Combustible.	No open flames. No contact with strong oxidants.	Alcohol-resistant foam, powder, water spray, carbon dioxide.
Explosion	Above 79°C explosive vapour/air mixtures may be formed.	Above 79°C use a closed system, ventilation.	In case of fire: keep drums etc. cool by spraying with water.
Inhalation	Sore throat. Burning sensation. Cough. Dizziness. Headache. Nausea. Vomiting. Shortness of breath. Laboured breathing. Unconsciousness.	Avoid inhalation of fine dust and mist. Ventilation or breathing protection.	Fresh air, rest. Half-upright position. Refer for medical attention.
Contact with skin	Easily absorbed. Serious skin burns. Numbness. Convulsion. Collapse. Coma. Death.	Protective gloves. Protective clothing.	Remove contaminated clothes. Rinse skin with plenty of water or shower. To remove substance use polyethylene glycol 300 or vegetable oil. Refer for medical attention. Wear protective gloves when administering first aid.
Contact with eyes	Pain. Redness. Permanent loss of vision. Severe deep burns.	Face shield, or eye protection in combination with breathing protection.	First rinse with plenty of water for several minutes (remove contact lenses if easily possible), then take to a doctor.
Ingestion	Corrosive. Abdominal pain. Convulsions. Diarrhoea. Shock or collapse. Sore throat. Smoky, greenish-dark urine.	Do not eat, drink, or smoke during work. Wash hands before eating.	Rinse mouth. Give plenty of water to drink. Do not induce vomiting. Refer for medical attention.
Spillage	n/a	n/a	Sweep spilled substance into sealable containers; if appropriate, moisten first to prevent dusting. Carefully collect remainder then remove to safe place. Personal protection: complete protective clothing including self-contained breathing apparatus. Do not let this chemical enter the environment.
Storage	n/a	Separated from strong oxidants, food and feedstuffs. Keep in a well-ventilated room.	n/a

Potassium bromide:

The hazards for this substance are almost negligible, especially for the low concentrations that I will be using. However, protective goggles will still be worn to avoid contact with eyes, as this could cause irritation. This substance could also cause some discomfort if ingested, so I will take be very careful to ensure that this doesn't happen.

Method for Investigating Temperature Change

Extra Apparatus	Reason For Use	Size/Concentration
Heating plate	Used to heat the mixture to different temperatures during the reaction.	n/a
Beakers	Used as a water bath to heat the mixtures evenly.	500ml
Boiling tubes	Used to hold reacting mixture whilst it is heated or cooled to the desired temperature.	n/a
Ice cubes	Used to cool the water bath to temperatures below room temperature.	n/a
Thermometer	Used to measure the temperature of the mixture so that it can be kept constant throughout the reaction.	n/a

Method

1. Set the apparatus up as was done in the experiment investigating concentration change. However, this time use a heating plate rather than a magnetic stirrer so that temperature can be adjusted.
2. Mix 5cm^3 of potassium bromate(V) solution, 5cm^3 of sulphuric acid solution and 5cm^3 of phenol solution in a boiling tube and add 4 drops of methyl orange indicator.
3. Add 5cm^3 of potassium bromide to a separate boiling tube.
4. Fill a large beaker with water and place it on top of the heating plate and set it to the required temperature. To achieve cooler temperatures put ice cubes into the beaker and monitor the temperature using a thermometer. Place the boiling tubes containing the correct solutions in the water bath so that they are also heated or cooled to this temperature.
5. Mix the two solutions and start the stopwatch. Record the time taken for the solution to go colourless. Ensure that the temperature of the water is kept constant. This can be done by having a thermometer in the beaker. Although the higher temperatures should be easily maintained by the heating apparatus, lower temperatures must be kept constant by adding more ice cubes if required.
6. Repeat 5 times with each temperature to ensure accurate and fair results.

Results With Respect to Potassium Bromate(V)

The table below shows my results when varying the concentration of potassium bromate(V):

Concentration of Potassium Bromate(V) (mol/dm ⁻³)	Time Take For the Mixture to Turn Colourless (seconds)						Reaction Rate (seconds ⁻¹)
	Repeat 1	Repeat 2	Repeat 3	Repeat 4	Repeat 5	Average	
0.01	31.9	32.2	31.7	31.8	32.0	31.92	0.0313
0.008	45.6	45.1	45.3	45.4	45.2	45.32	0.0221
0.006	62.8	63.4	62.7	62.9	63.2	63.00	0.0159
0.005	88.8	88.3	89.0	88.5	88.7	88.66	0.0113
0.004	104.5	105.2	105.6	104.9	105.2	105.08	0.0095
0.003	121.3	122.1	121.9	121.7	122.0	121.80	0.0082

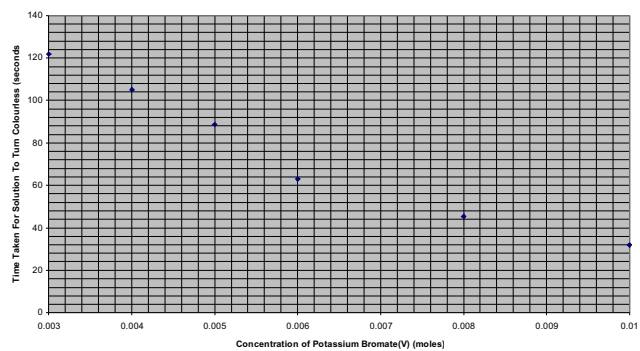
Graph 1 on the next page shows the concentration of potassium bromate(V) plotted against the average time taken for the solution to turn colourless. From this graph I can see that there is a negative correlation because as concentration is increased, the time taken for the solution to turn colourless decreases. However, this graph does not provide enough information to work out the rate equation data for potassium bromate(V). For this, I had to draw up a graph of concentration against reaction rate. Reaction rate is worked out using the following equation:

$$\text{Reaction Rate} = \frac{1}{\text{Time}}$$

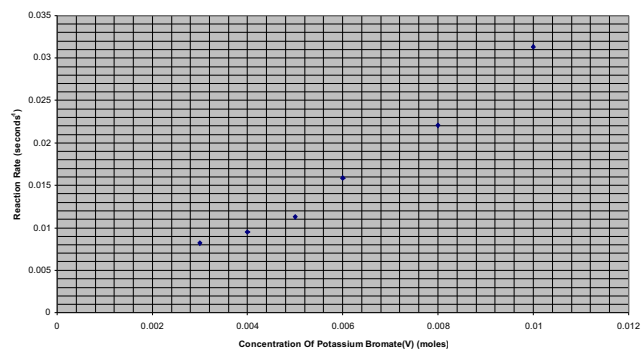
This is presented in Graph 2. Graph 2 has a line of best fit which is a straight diagonal line, rather than a curve. It also shows that there is a positive correlation between concentration and reaction rate, as when concentration of potassium bromate(V) is increased, reaction rate increases. I could draw the line of best through the origin at point (0,0) as I know that if the concentration is 0M then the reaction will not occur – the reaction rate will be 0seconds⁻¹ (this can be applied to all other graphs of concentration against reaction rate that I have drawn). I decided to draw the majority of my graphs using a computer to plot the axes and points, and drawing the line or curve of best fit by hand. The computer can provide better accuracy than I can as far as drawing the axes and plotting the points is concerned. However, I believe that I can provide more accuracy when drawing the line or curve of best fit. My line of best fit suggests that the reaction, with respect to potassium bromate(V), is first order. This means that that rate equation so far is as follows:

$$\text{Rate} = k[\text{BrO}_3]$$

Graph To Show Concentration Of Potassium Bromate(V) Against Time



Graph To Show Concentration Of Potassium Bromate(V) Against Reaction Rate



Results With Respect to Potassium Bromide

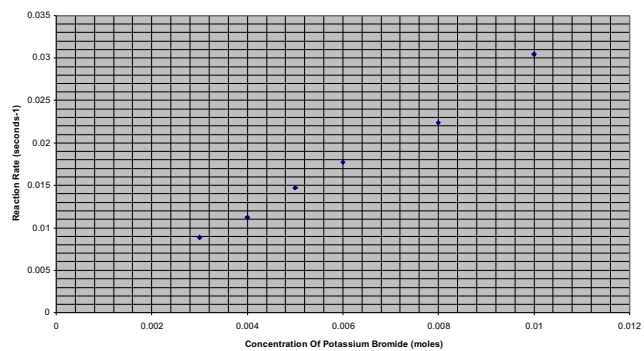
The table below shows my results when varying the concentration of potassium bromide:

Concentration of Potassium Bromide (mol/dm ⁻³)	Time Take For the Mixture to Turn Colourless (seconds)						Reaction Rate (seconds ⁻¹)
	Repeat 1	Repeat 2	Repeat 3	Repeat 4	Repeat 5	Average	
0.01	33.2	32.8	33.1	32.5	32.7	32.86	0.0304
0.008	44.7	44.2	45.2	44.5	44.8	44.68	0.0224
0.006	56.1	56.5	56.9	56.2	56.4	56.42	0.0177
0.005	68.3	67.9	67.8	68.1	67.7	67.96	0.0147
0.004	89.6	89	88.9	89.3	89.4	89.24	0.0112
0.003	112.9	113.5	112.7	113.2	113.3	113.12	0.0088

Graph 3 on the next page shows the concentration of potassium bromide plotted against reaction rate. The line of best fit is once again a straight diagonal one. This shows that there is again a positive correlation between concentration and reaction rate, as when the concentration of potassium bromide is increased, the reaction rate increases. This means that the reaction, with respect to potassium bromide, is also first order. This in turn means that the rate equation so far looks like this:

$$\text{Rate} = k[\text{BrO}_3][\text{Br}^-]$$

Graph To Show Concentration Of Potassium Bromide Against Reaction Rate



Results With Respect to Sulphuric Acid

The table below shows my results when varying the concentration of sulphuric acid:

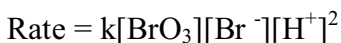
Concentration of Sulphuric Acid (mol/dm ⁻³)	Time Take For the Mixture to Turn Colourless (seconds)						Reaction Rate (seconds ⁻¹)
	Repeat 1	Repeat 2	Repeat 3	Repeat 4	Repeat 5	Average	
1	28.4	29.1	28.7	28.6	28.9	28.74	0.0348
0.8	44.3	45.1	45.2	44.5	44.8	44.78	0.0223
0.6	101.7*	64.9	65.4	64.8	65.0	65.03	0.0154
0.5	98.9	99.3	99.5	98.8	99.4	99.18	0.0101
0.4	Reaction Took Too Long						
0.3	Reaction Took Too Long						

As can be seen from the table above, at concentrations of 0.3 and 0.4M, the solution took much too long to turn colourless. This was impractical, as it meant that I was less likely to be able to conduct a suitable number of repeats to get an accurate average time. Because of this I was forced to make some minor modifications to my original method. As I only had a range of four concentrations recorded for sulphuric acid, I knew that this would not be enough to draw a very useful graph of concentration against reaction rate. For this reason, I decided to conduct experiments testing two more concentrations within the range of the other four, thus making up for the two concentrations I had not been able to investigate. The results for these extra concentrations are shown below:

Concentration of Sulphuric Acid (mol/dm ⁻³)	Time Take For the Mixture to Turn Colourless (seconds)						Reaction Rate (seconds ⁻¹)
	Repeat 1	Repeat 2	Repeat 3	Repeat 4	Repeat 5	Average	
0.9	36.5	36.1	36.9	36.7	36.4	36.52	0.0274
0.7	54.3	54.5	53.9	54.2	54.1	54.20	0.0185

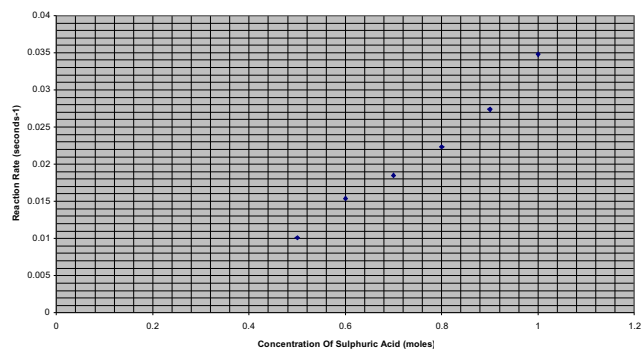
The result in my table highlighted both in bold and by an asterisk is clearly an anomaly. For this reason I decided to leave it out when calculating the reaction rate at 0.6M. Although this meant that I only used four repeats for this concentration, the average results I obtained still coincides with other results and is therefore still relatively accurate.

Graph 4 on the following page shows the concentration of sulphuric acid plotted against the reaction rate. From my line of best fit it is clear that there is a positive correlation between the two. My line of best fit is in a curved shape. Relating back to my theory section, this shows that the reaction, with respect to sulphuric acid, is second order. This in turn means that I have proved the final rate equation to be:



This is exactly what I predicted at the beginning of my investigation.

Graph To Show Concentration Of Sulphuric Acid Against Reaction Rate



Now that I have proved the rate equation experimentally, I want to find a value for the constant, 'k'. This is called the rate constant ^(1, page 229) and its value changes with temperature. Using my results at a certain temperature I should be able to calculate a value for 'k'. I will use the following values:

$$[\text{BrO}_3] = 0.008$$

$$[\text{Br}^-] = 0.01$$

$$[\text{H}^+]^2 = (0.1)^2 = 0.01$$

At room temperature, I calculated the reaction rate for these values to be $0.0221 \text{ seconds}^{-1}$. I will now substitute the above values into the rate equation in order to calculate the constant 'k', although I first need to rearrange the equation to make 'k' the subject:

$$\text{Reaction Rate} = k[\text{BrO}_3][\text{Br}^-][\text{H}^+]^2$$

$$k = \frac{\text{Reaction Rate}}{[\text{BrO}_3][\text{Br}^-][\text{H}^+]^2}$$

$$k = \frac{0.0221}{(0.008)(0.01)(0.01)}$$

$$k = \frac{0.0221}{(8 \times 10^{-7})}$$

$$k = 27625$$

The units for 'k' vary depending on the reaction investigated. In this case they are worked out as follows:

$$k = \frac{(\text{mol dm}^{-3})(\text{seconds}^{-1})}{(\text{mol dm}^{-3})(\text{mol dm}^{-3})(\text{mol dm}^{-3})^2}$$

$$k = 27625 \text{ s}^{-1} \text{ mol}^{-3} \text{ dm}^9$$

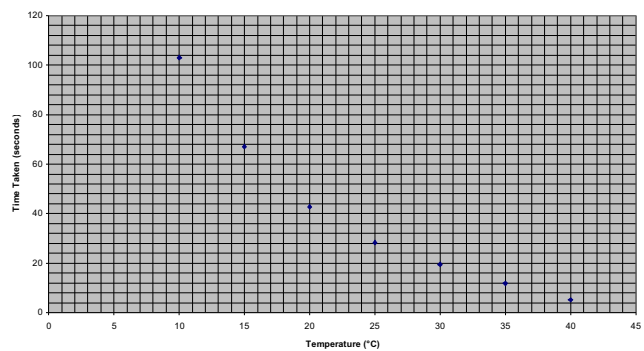
As I mentioned earlier, a change in temperature alters the value for 'k'. The value I calculated above was taken at room temperature. On the following page is a table showing my results when I investigated how temperature affects the reaction rate.

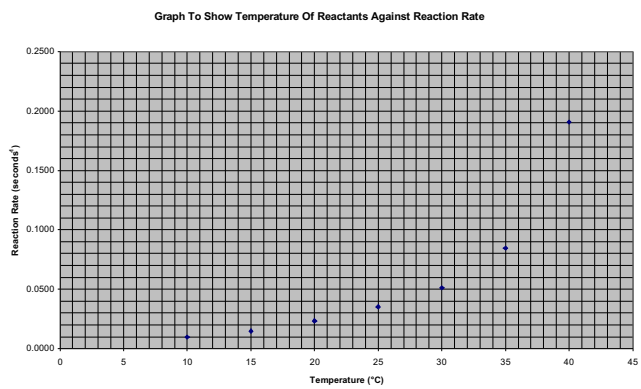
Results When Varying Temperature

Temperature of Reactants (°C)	Time Take For the Mixture to Turn Colourless (seconds)						Reaction Rate (seconds ⁻¹)
	<i>Repeat 1</i>	<i>Repeat 2</i>	<i>Repeat 3</i>	<i>Repeat 4</i>	<i>Repeat 5</i>	<i>Average</i>	
10	102.7	103.1	103.2	102.9	103	102.98	0.0097
15	67.2	67.4	66.6	67.1	66.8	67.02	0.0149
20	42.4	42.9	43.1	42.6	42.8	42.76	0.0234
25	28.8	28.3	27.9	28.1	28.6	28.34	0.0353
30	19.6	19.8	19.3	19.5	19.4	19.52	0.0512
35	12.1	11.5	11.9	11.7	11.9	11.82	0.0846
40	5.1	5.6	4.9	5.2	5.4	5.24	0.1908

From Graph 5 on the next page, it is clear that there is a negative correlation between temperature and time taken for the solution to turn colourless – as the temperature of the reactants is increased, the time taken for the solution to turn colourless decreases. Graph 6 shows that there is a positive correlation between temperature and reaction rate – as the temperature of the reactants is increased, the reaction rate increases. Both graphs have a curve of best fit, rather than a straight diagonal line.

Graph To Show Temperature Of Reactants Against Time Taken For Solution To Turn Colourless





Graph 6

Using the rearranged Arrhenius equation discussed in my theory, I can draw up a table of values needed to work out the activation enthalpy for this reaction:

Temperature of Reactants (°C)	Temperature of Reactants (K)	1/Temperature (K ⁻¹)	Reaction Rate (s ⁻¹)	Log To The Base e Of Reaction Rate
10	283	0.00353	0.0097	-4.636
15	288	0.00347	0.0149	-4.206
20	293	0.00341	0.0234	-3.755
25	298	0.00336	0.0353	-3.344
30	303	0.00330	0.0512	-2.972
35	308	0.00325	0.0846	-2.469
40	313	0.00319	0.1908	-1.657

These results are plotted in Graph 7 on separate graph paper. This graph shows 1/temperature in Kelvin against log to the base e of reaction rate. I chose to use a completely hand-drawn set of axes so that I could accurately work out the gradient of the line, thus enabling me to work out an accurate value for the activation enthalpy. I could not easily draw the triangle on the line of best fit, which allows me to calculate the gradient, using a computer. Graph 7 shows that there is a negative correlation between the two sets of data – as 1/temperature is increased, log to the base e of reaction rate decreases.

To find the gradient of a straight line, you use the following equation: ⁽⁷⁾

$$\text{Gradient} = \frac{\text{Change in y value}}{\text{Change in x value}}$$

For my graph, I worked out the gradient as follows:

$$\text{Gradient} = \frac{-1.75}{0.00024}$$

$$= -7291.7$$

$$\text{Gradient} = -E_a = \frac{-7291.7}{R}$$

$$R \text{ (the gas constant)} = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$$

The gradient equation can be rearranged as:

$$-E_a = \text{gradient} \times R$$

$$= -7291.7 \times 8.31$$

$$= -60594$$

$$E_a = 60594 \text{ J mol}^{-1}$$

This can be converted to kJ mol^{-1} by dividing by 1000 (the number of joules in a kilojoule):

$$\frac{60594}{1000} = 60.594 \text{ kJ mol}^{-1}$$

This is the activation enthalpy for my reaction.

Conclusion

I have already drawn out all of my results tables and have drawn graphs for each one. From these I have been able to work out the rate equation, rate constant and activation enthalpy for this reaction. Therefore I have fulfilled all of the aims that I set out to complete.

I collected four different sets of results throughout my experiment. The first three sets of results concerned the concentration of different substances in the reaction. I varied these to try and find the rate equation. In general, I found that the solution turned colourless at a much slower speed at lower concentrations. At the same time this means that the reaction occurred quicker at higher concentrations. This can be simply explained by referring back to the collision theory that I mentioned earlier in my investigation. This suggests that when there is less reactant molecules it is less likely that there will be a collision between two particles. If molecules are not colliding as often then reactions between molecules will also occur less often. This means that the reaction occurs at a much slower rate. Graph 1 shows this relationship in my investigation - as the concentration of potassium bromate(V) is decreased, the amount of time taken for the solution to turn colourless, and thus the reaction to finish, increases. This is a strong negative correlation.

The reaction rate of a reaction can be worked out using the following equation:

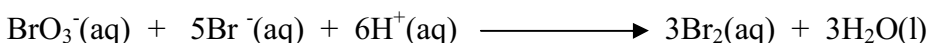
$$\text{Reaction Rate} = \frac{1}{\text{Reaction Time}}$$

I calculated the reaction rate for all of my sets of results and then plotted graphs showing concentration of solution against reaction rate. I could then determine the order of reaction for each substance used in the reaction. I already knew from my research that the reaction should be first order, with respect to potassium bromate(V) and potassium bromide, and second order with respect to sulphuric acid. I was able to clearly prove the correct order of reaction for each of these chemicals by using my graphs of concentration against reaction rate. To work out the overall order of the reaction all of the individual orders of reaction must be added together. As I have correctly calculated the individual orders of reaction, I can accurately state that the overall order of reaction is as follows:

$$\begin{aligned}\text{Overall order} &= 1 (\text{KBrO}_3) + 1 (\text{KBr}^-) + 2 (\text{H}_2\text{SO}_4) \\ &= 4\end{aligned}$$

Every reaction has a reaction mechanism. This looks at the individual 'steps' in the reaction which may not obviously be occurring. Using the rate equation I can work out how many of each ion is involved in the rate determining step. The rate determining step is described as the slowest step in the reaction mechanism. To find how many ions are used in this step, you simply look at the power to which each ion is raised. For example, the reaction, with respect to H^+ ions, is second order. This means that there are two H^+

ions involved in the rate determining step. The general equation for this reaction is as follows:

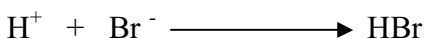


The rate equation, as previously proved, is:

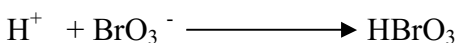
$$\text{Rate} = k[\text{BrO}_3^-][\text{Br}^-][\text{H}^+]^2$$

I can now see that the rate determining step involves one BrO_3^- ion, one Br^- ion and two H^+ ions. It is likely that this slow rate determining step is preceded by faster reactions. The mechanism which I suggest occurs is as follows:

Step 1: (fast)

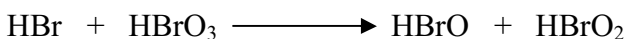


Step 2: (fast)



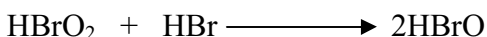
The HBr and HBrO_3 formed in the previous two reactions now react in Step 3.

Step 3: (rate determining step)

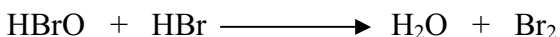


The products of Step 3 then react further to give water and bromine in Step 4 and 5.

Step 4: (fast)



Step 5: (fast)



Reactions mechanisms cannot be proven – they are just suggestions to what happens in reactions, based on a good understanding of chemistry. The reaction mechanism above is simply a suggestion as to what may happen in this reaction.

Concentration was not the only thing that I varied in my investigation. I also chose to vary the temperature of the reactants. I already knew from my research that changing temperature has an affect on rate of reaction, as it changes the value of the rate constant, 'k'. My results and graphs show clear correlation that if temperature is increased, reaction rate increases. This means that the time taken for the solution to turn colourless decreases

as temperature is increased. I also found that even a small increase in temperature could dramatically affect the reaction rate. For example, when I changed the temperature of the reactants from 10 to 20°C, the average time taken for the reaction to occur fell from 102.98 to 42.76 seconds. This means that, with a temperature change of just 10°C, the time take for the solution to turn colourless fell by 60.22 seconds.

The reason for this dramatic change in reaction rate can again be explained by the collision theory and kinetics. As temperature increases, molecules gain more kinetic energy and can therefore move quicker. The fact that they are moving more quickly means that molecules are statistically more likely to collide. This causes an increase in reaction rate. Molecules will not only collide more frequently with higher temperatures, but they will collide more often with the required energy for a reaction to occur. This required energy is called the activation enthalpy. As temperature is raised, more molecules reach the activation enthalpy and therefore the reaction is more likely to take place sooner.

I plotted a graph of log to the base e of reaction rate against the reciprocal of the temperature in Kelvin (1/temperature). This gave me a gradient of $-E_a/R$ as is mentioned in the section earlier in my investigation about the Arrhenius equation. Using this graph I could then work out the activation enthalpy. I found it to be $60.594 \text{ kJ mol}^{-1}$. The Maxwell-Boltzmann diagram shown in my theory section explains this.

Earlier in my investigation I predicted that with an decrease of just 10°C, between 20 and 10°C, the reaction time would at least double. The percentage increase for my results can be worked out as follows:

Reaction time at 10°C = 102.98 seconds

Reaction time at 20°C = 42.76 seconds

$$42.76 \times ? = 102.98$$

$$? = \frac{102.98}{42.76}$$

$$? = 2.41$$

This means that, in my experiment, a decrease of 10°C brought about a 241% increase in reaction time. This shows that the effects of temperature are very important on reaction rate.

At room temperature, I calculated the value for the rate constant, 'k'. I worked out a value of $27625 \text{ s}^{-1} \text{ mol}^{-3} \text{ dm}^9$. Although the value for 'k' does not change with concentration, it is affected by a change in temperature. This is why the value for rate of reaction changes as temperature is changed.

In my reaction an orange solution turns colourless after time. This is because the bromine that is being formed, as shown in my suggested reaction mechanism, reacts with phenol. Once the bromine has filled all of the binding sites on the phenol present in the solution, it reacts with the methyl orange indicator. This causes the orange colour to turn colourless.

When I was varying temperature, I obtained an anomalous result. This result clearly did not fit in with the trend that was developing with the rest of the results. I chose to ignore this result as it would obscure my average reaction time and therefore my calculations for rate of reaction would also be affected. I will elaborate on this and other problems in my evaluation section.

Evaluation

In general I think that my investigation went well. I succeeded in all of my initial aims and produced accurate, reliable results. I only encountered one real anomaly throughout the investigation process (highlighted in my conclusion).

I conducted a pre-test in my planning section of my investigation. This allowed me to highlight any problems with my initial method and I therefore had a chance to rectify them, rather than encountering problems later on. I investigated suitable concentrations of each solution to use. I noticed that at much lower concentrations, the reaction went much too slowly. If the reaction goes too slowly I would not have time to do five repeats for each concentration. This would reduce both accuracy and reliability of my results. For this reason I decided to exclude the two lowest concentrations of each solution. Although this meant that I had lost some of the range of my results, I still investigated six different concentrations for each substance. This means that I was still able to obtain accurate, reliable results.

Another thing I wanted to find out in my pre-test was how to stir the solution whilst the reaction was taking place. I conducted an experiment to investigate this. I found that the most useful method of stirring would be to use a magnetic stirring device. This ensured that I only had to concentrate on one thing – the point at which the solution turned completely colourless. If I had to stir the solution by hand I may not notice the colour change straight away. Using a magnetic stirrer I could ensure that my results were more accurate. I also put my conical flask, with reacting solution in it, on a white tile. This meant that any colour change was more obvious, as it was against a white background. The surface of the table I was working at was a darker colour, meaning that if the conical flask was placed straight onto the table it would be much harder to see when the solution had turned colourless.

I also conducted a pre-test to trial my method for investigating how temperature change affects the reaction rate. I had originally decided to monitor only the temperature of the water bath heating my solutions in boiling tubes. However, I realised that this would most likely not share the same value as the temperature of the solutions in the boiling tubes. To overcome this problem I placed a thermometer in each boiling tube – one in the tube containing potassium bromide and one in the solution of the other reactants. Once both had reached the correct temperature I was able to mix the two solutions, knowing that both were at the required temperature. I also discovered that when investigating temperatures above 40°C the reaction became too fast to record accurate timings. I would have lowered the concentration of potassium bromide, although I wanted to keep this constant to my previous experiments. This would mean that I could calculate a more accurate rate equation. Instead, I chose the following range of temperatures: 10, 15, 20, 25, 30, 35 and 40°C. This range made it easier for me to keep the temperature of the reactants stable, as they do not fluctuate to far above or below room temperature.

I took into account how precise and reliable all of my separate measurements throughout the investigation were. This is why I decided to use burettes instead of measuring

cylinder. As explained in my pre-test section, the percentage error associated with measuring 5cm^3 of solution in a burette is just 2%. This is much more accurate than a measuring cylinder. The percentage error for a measuring cylinder is as follows:

$$\text{Percentage error} = \frac{\text{error}}{\text{measured value}} \times 100$$

So for measuring, for example, 5cm^3 from a measuring cylinder (with error $\pm 0.5\text{cm}^3$) the error would be:

$$\text{Percentage error} = \frac{0.5}{5} \times 100$$

$$= 10\%$$

This is a much larger percentage error than for a burette, so I made the right decision to ensure accuracy by using a burette rather than a measuring cylinder.

The volumes of solution I used were as follows:

Potassium bromate(V)	$= 5\text{cm}^3$
Potassium bromide	$= 5\text{cm}^3$
Phenol	$= 5\text{cm}^3$
Sulphuric acid	$= 5\text{cm}^3$

The error for all of these measurements is just 2%. This small percentage error is almost negligible in terms of how much it will have affected my results, due to the fact that the error is so small. I could also use all of these volumes when varying temperature.

Although I was now using boiling tubes rather than a conical flask to house the reactions, the fact that all volumes were just 5cm^3 meant that none of the volumes had to be changed in order to fit all the solutions into a single boiling tube.

In my experiment I chose to study rate of reaction as a 'clock reaction'. This means that rate is measured by timing how long it takes to produce a certain amount of one product. In my experiment this product was bromine. If the rate of reaction is high then the reaction time must be lower, as rate is the inverse of reaction time. This is shown as:

$$\text{Rate} = \frac{1}{\text{Reaction Time}}$$

This method works in my reaction due to the presence of phenol. At first any bromine formed reacts with phenol. However, when all of the binding sites on all of the phenol molecules present are filled, bromine begins to react with the methyl orange indicator. This in turn causes the solution to turn colourless by decolourising the methyl orange indicator. Because I kept the concentration and amount of phenol, and the amount of methyl orange added, the same throughout my investigation I know that the amount of

bromine needed in each test is the same. This ensures that the test is fair. This is an accurate method as it is simple and therefore it is easy to avoid mistakes. I did not need to take a series of measurements throughout the reaction, which increases the chances of human error. By plotting a graph of reaction rate against concentration I could easily find the order of reaction. I also did not need to draw tangents to any of my curves, which further reduces any chance of human error when drawing and analysing my data in graph form.

I tested four variables in my investigation. The first variable I tested was the concentration of potassium bromate(V). I drew up results tables and drew a graph of reaction rate against concentration of potassium bromate(V). I found that the reaction was first order with respect to potassium bromate(V). This was exactly as I had predicted in my aim. I did not encounter any anomalous results whilst investigating this variable and so can be sure that these results are accurate. I carried out five repeats for each concentration for each variable, ensuring that I had accurate and reliable results.

Next I investigated how changing the concentration of potassium bromide affected the rate of reaction. Again I drew up tables and a graph showing reaction rate against concentration of potassium bromide. I found that the reaction was also first order with respect to potassium bromide. This again was as I had predicted earlier on in my investigation. There were no anomalous results for potassium bromide, so I am sure that my results are accurate.

I conducted an experiment to test how the concentration of sulphuric acid affects the rate of reaction. I drew up a results table and drew a graph of reaction rate against concentration of sulphuric acid. I found that the reaction was second order with respect to sulphuric acid. This was what I predicted in my aim. However, I did have an anomalous result whilst recording my results for sulphuric acid. This came at a concentration of 0.06M and the value was 101.7seconds. When compared with the other results for the same concentration (65.4, 64.9, 64.8 and 65.0seconds) this is clearly an anomaly. Although this could have potentially affected my values for rate of reaction, I came up with a solution. Because I had conducted five repeats for each concentration, I knew that the exclusion of one of these repeats would not have a huge affect on my results. For this reason I was able to exclude the value of 101.7seconds when calculating an average reaction time for the concentration of 0.06M. This also meant that it was not included when further investigating the order of reaction for sulphuric acid. The fact that I still found that the order of reaction, respect to sulphuric acid, was two means that my results were still accurate even when excluding this anomaly.

The aforementioned anomaly could be caused by several different things. My own human error may have meant that I produced a solution of sulphuric acid with a lower concentration than I had intended to. This would make the reaction go much slower than in the other repeats. One other possibility is that the solution had somehow become contaminated. If another chemical was included in the sulphuric acid solution, it would change how quickly the reaction occurred. This could be due to not washing out the apparatus thoroughly enough. Another possible reason is simply that I did not notice the

colour change at the right time. This is human error but I don't think that this was very likely, as the colour change is very easy to see. I think that the most likely reason was making up the solution incorrectly, as this is most likely to affect the reaction time as greatly as the anomaly suggests. For this reason I disposed of the sulphuric acid solution and made up a fresh batch. This appears to have worked, as in my results table for sulphuric acid the values were much closer together and continued the trend, after the first repeat had been ignored. This human error can easily be avoided by ensuring that each solution is measured out very carefully.

Another potential problem that I could have encountered would have been if I had added too much phenol, or had added phenol of a stronger concentration. This would mean that there were more phenol molecules. If there are more phenol molecules then the bromine made in the reaction may not exceed the amount of phenol. This would in turn mean that all of the phenol binding sites are not filled, and bromine does not get a chance to react with the methyl orange indicator. This would mean that the solution would not turn colourless.

Overall, I think that my investigation was a success. I managed to prove that the rate equation for the bromine clock reaction was as follows:

$$\text{Rate} = k[\text{BrO}_3^-][\text{Br}^-][\text{H}^+]^2$$

This is exactly as I had predicted in my aim.

I investigated the affect that temperature has on the rate of reaction, and found that the rate constant, 'k', changes with temperature. I then used this knowledge, along with the Arrhenius equation to find the activation enthalpy of this reaction. I found that the activation enthalpy in my investigation was 60.594kJmol^{-1} . This seems to be a reasonable value, and I can therefore conclude that I achieved all of my aims.

References

¹ Salters Advanced Chemistry - Chemical Ideas (Heinemann) - page numbers included where referenced.

² <http://www.science.uwaterloo.ca/~cchieh/cact/c123/coneffec.html>

³ <http://www.chemguide.co.uk/physical/basicrates/temperature.html>

⁴ <http://www.chemguide.co.uk/physical/basicrates/arrhenius.html>

⁵ <http://www.hometrainingtools.com/articles/making-chemical-solutions-science-teaching-tip.html>

⁶ <http://www.inchem.org>

⁷ http://www.mathsteacher.com.au/year10/ch03_linear_graphs/02_gradient/line.htm