The Effects of Strong and Weak Acids.

Aim

To investigate the effect strong and weak acids have on the rate of a reaction and its order.

Introduction

What is a chemical reaction? This is the first question that needs to be answered before tackling this investigation. A chemical reaction is the process which results in the chemical conversion of one substance into another. These reactions can be placed into two groups; exothermic and endothermic reactions. If the energy given out of the reaction is greater than the energy taken in, the reaction is known as an 'Exothermic Reaction'. Similarly, if the energy taken in is greater than the energy given out, the reaction is known as an 'Endothermic Reaction'. The easiest way to determine if a reaction is Exo/Endothermic is to observe any temperature changes within or after the reaction has taken place. If the surrounding temperature drop, the reaction is exothermic and if the temperature increases, the reaction is exothermic.

Although it is true that every substance has the potential to be converted into another, it is also true that before a reaction happens, certain criteria needs to be met. For a reaction to happen, the particles must collide with the correct amount of energy – also known as the 'Activation Energy'. The Activation energy is the lowest amount of energy a particle needs in order for it to successfully collide with another and for the reaction to take place. So, for a reaction to 'go', a certain number of particles have to have the correct amount of energy. Although the activation energy for a reaction can never be changed, outside factors can affect how many particles in the reaction have the minimum energy requirement to react successfully.
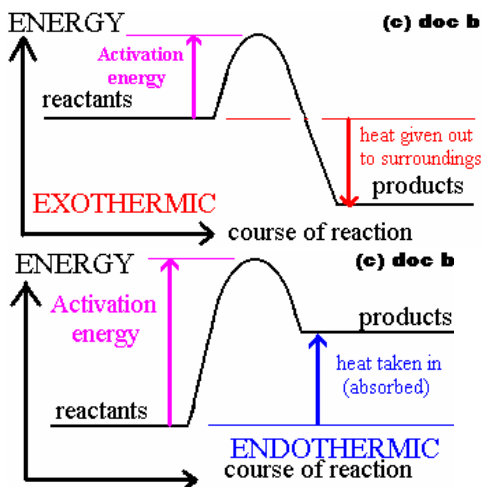

Figure 1: http://www.docbrown.info/page03/3_31rates.htm#3a

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As you can see from the above diagrams, the Activation energy (Ea) is the difference between the transition state of the particles – the state where the old bonds aren't yet fully broken and the new bonds aren't yet fully formed.

You will also notice that Endothermic reactions have a much greater Activation energy than Exothermic reactions and this is because the energy taken to form new bonds is greater than the energy taken to break them.

The rate of a reaction is the time taken for the particles to reach the activation energy and for the reaction to 'go'. The fewer the number of particles with the activation energy, the slower the rate of reaction, and vice-versa. Not every particle in a substance can have the minimum energy requirement for the reaction and this is when outside factors can affect the rate of the reaction. Outside factors, such as the temperature, concentration, surface area and the use of a catalyst can affect the rate of the reaction by speeding it up or slowing it down.

Effect of Surface Area

The number of correct collisions in the reaction mixture affects the rate of the reaction. This is directly proportional to the surface area of the reactants. The greater the surface area, the higher the chance of correct collisions taking place. The smaller the surface area, the smaller the chance of correct collisions taking place. This is due to the number of particles exposed to the others – the greatest example of this is the use of Magnesium Powder in comparison to Magnesium Ribbon.
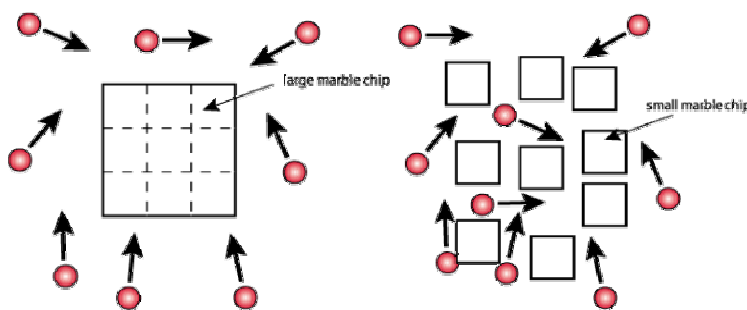

Figure 2:
http://www.sciencepages.co.uk/keystage4/GCSEChemistry/m3ratesofreaction.php#rate7

As you can see, the smaller chips – or Mg powder – would be more exposed to the reactants and will therefore encounter more successful collisions. This will increase the rate of reaction. However, sometimes the particles can be too small and increase the rate of reaction too much for it to be observed. It is for this reason that Mg ribbon would be the easiest to use in my experiment – more on that later on.

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Effect of Concentration

The increase in the concentration of a substance will increase the rate of its reaction. This is because there are more particles to be reacted with and therefore more collisions to be made – this will increase the number of successful collision which is what makes the reaction 'go'. The concentration of something is the number of moles of a product per given volume. SO the increase in concentration means that more particles are in closer contact with each other and the collision between the correct molecules is more likely.
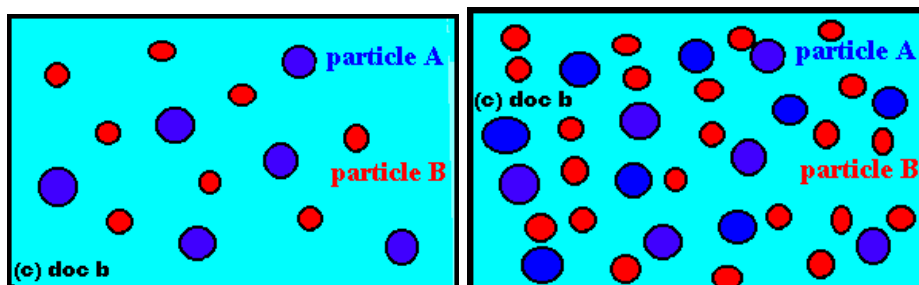

Figure 3: http://www.docbrown.info/page03/3_31rates.htm#3a

Effect of temperature

As stated previously, every particle has some energy within them – this is called Kinetic energy. As the temperature is increased, the kinetic energy these particles have is also increased. This means that the particles are not only moving faster – which will help them surpass the activation energy – but there is also a greater chance of correct collisions. With more particles able to react now due to their increase in kinetic energy, the number of correct collisions will increase which means the overall rate of reaction will increase.
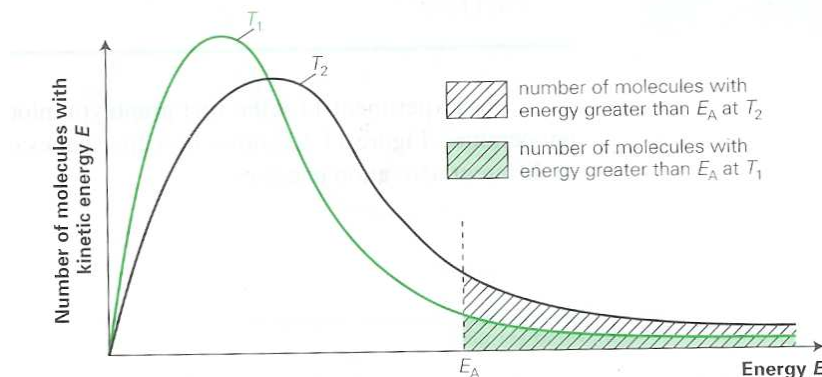
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As you can see from the diagram above, T_1 is at a lower temperature than T_2 and has a significantly lower number of molecules with energy greater than the Activation energy.

Effect of a Catalyst

The use of a catalyst increases the rate if a reaction. It is wrong to say, however, that the catalyst lowers the activation energy of a reaction. The definition of a catalyst is a substance which provides an alternate route for the reaction with a lower activation energy and can be used in a reaction without getting used up, itself. By providing an alternate route for the particles with a lower activation energy, more particles have the minimum energy requirement and so can take part in the reaction – this leaves more scope for the number of correct collisions per second.
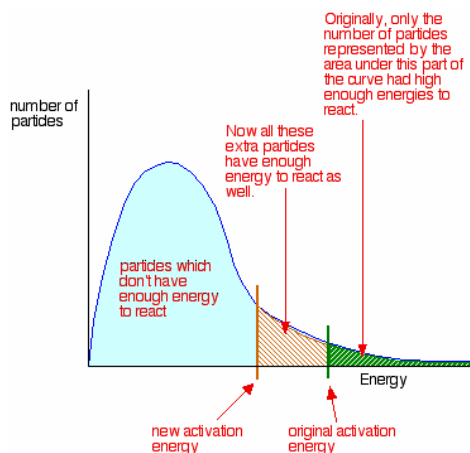

Figure 4: http://www.chemguide.co.uk/physical/basicrates/catalyst.html

As you can see from the figure above, the use of a catalyst shows that a significantly larger proportion of particles have the minimum needed amount of energy to take part in the reaction.

While an alternate activation energy is provided, the thermodynamics of the reaction remain unchanged. There are two types of catalysis – Homogenous and Heterogeneous.

Homogenous catalysis is when both the reactants and the catalyst are in the same phase. An example of this would be like the synthesis of an Ester; both the Carboxylic acid, the alcohol and the Hydrogen ion catalyst used are in an aqueous solution.

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Heterogeneous catalysis is when both the reactants and the catalyst are in different phases. An example of this would be in the Haber process in the synthesis of Ammonia. The reactants (Hydrogen and Nitrogen) are gases and the catalyst (Iron) is a solid.

<u>Working out the Activation Energy (Ea)</u>

To work out the activation energy of a reaction, the Arrhenious equation must be used. This is:

$$\text{rate} = Ae^{-(Ea \ / \ RT)}$$

Where, A = constant,

e = exponential function,

R = gas constant and

T = temperature.

This equation can also be written as:

$$\ln (\text{rate}) = \ln (A) -(Ea \ / \ RT)$$

ln, being the natural log of the number. Using the natural log makes the question far easier to use.

To work out the activation energy, one can substitute the equation above and write it as:

$$y = mx + c$$

This is where, y = ln (rate),

m (the gradient) = $-E_A/R$,

x = 1/T and

c = ln (A)

Once the gradient of the graph has been determined, one can then begin to substitute the numbers into the equation: -Ea = -gradient/8.31; the overall Activation energy number becoming positive sure to the negative signs in the equation.

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Working out the Rate of a Reaction

The rate of a reaction can be determined by doing the 'Rate Equation'. This is an equation which has gained by doing experiments and finding out the ratios between the concentrations of the substances.

$$r_A = k\ [A]^a[B]^b[C]^c$$

Where, Ra is the measure of the rate of reaction in $moldm^{-3}s^{-1}$,

k is the constant,

[A][B][C] are the concentrations of the substances in $moldm^{-3}$

And abc are the orders of the reaction based on those particular substances. To work out the overall order of the reaction, one needs to just add up these numbers.

Depending on the number of concentrations above, the units of Ra vary.

The Order

After doing the necessary experiments, values would have been obtained. To work out the order of a certain reaction with respect to another substance, one has to draw a time / Vf-Vt graph. By drawing a curve or a line, the order of the reaction can be obtained.
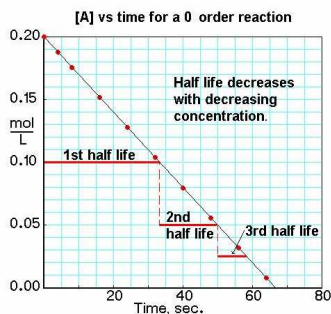
If the graph is just a straight line, it means the gradient has remained a constant and therefore the reaction has Zero order with respect to that particular substance. This is because the gradient constant shows that any change of concentration will not affect it.
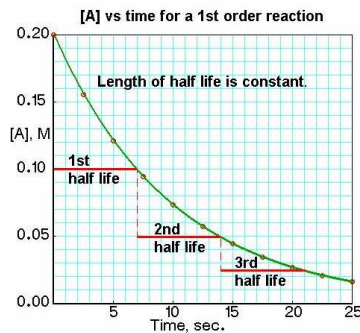
With curved graphs, one needs to determine the successive half lives before being able to determine the order of the reaction.

For First order graphs, the successive half lives are equal. This means that the time taken for the concentration of the reactants to be halved is constant.
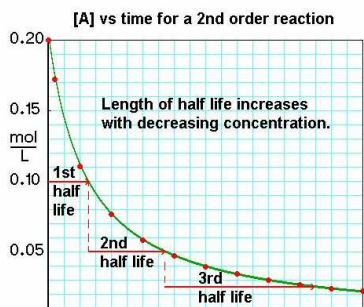
For Second order graphs, the successive half lives of the curve increase in size. This shows that the time taken for the concentration of the

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reactants to be halved *is*/*is* constant. It is also important to note that Second order graphs tend to be deeper than first order graphs, however half-life comparisons must be done in order to gain an accurate reading of the two.

<u>Following the Rate of a Reaction</u>

If there is any significant change shown within the reaction, it can be measured by using one or more of the following methods:

Dilatometry – When there is any significant changes made to the volumes of the liquids used in the reaction, this can be followed using a dilometer. This is a capillary tube with a scale on it so that any changes made to the volumes can be easily recorded over time.

Titration – A sample of the reaction mixture would be taken out at separate time intervals and quenched (to stop any reaction from taking place) by placing them in ice water. It is then titrated, using a burette, with a known concentration of another reagent. An indicator will be added to the mixture and any colour changes within the indicator will be observed.

Colorimetry – If any of the reactants or products has an intense colour, this can observed at different time intervals to measure the intensity of the colour. This can be followed by using a Colorimeter.

Electrical Conductivity – The electrical conductivity of a solution changes as the numbers of ions do. The total number of changed ions can be followed by measuring the electrical conductivity of the solution. This can be done by using a conductivity meter.

Gaseous Product – If the reaction gives of any gas, it can be collected in a gas syringe and measured at constant time intervals. By calculating the amount of gas collected over the set amount of time, the rate of the equation can be determined.

<u>Acids</u>

The definition of an acid is a substance which will donate a proton. This definition was made by Brønsted–Lowry and was an adjustment of the Arrhenious definition - 'An acid is a compound that will form hydrogen ions in water and contains hydrogen'. An example of the Brønsted–Lowry theory can be seen in the reaction between Water and Hydrogen Chloride gas:

$$HCl_{(g)} + H_2O_{(l)} \rightarrow Cl^-_{(aq)} + H_3O^+_{(aq)}$$
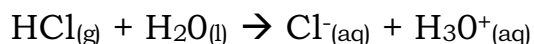
In this reaction, the Hydrogen is taken from the Chlorine and instead joins onto the water molecule to form Oxonium Ion.

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There are many different types of acids available. There can be strong and weak acids and monobasic and dibasic acids. A strong acid is defined as being totally disassociated in water. This means that the majority of acidic particles have been transferred in the reaction. An example of a strong acid would be Hydrochloric Acid which has almost 100% of its ions disassociated in water. A weak acid is defined as one that is only partially disassociated in water. An example of a weak acid would be Ethanoic Acid as only a small percentage of its ions can be disassociated in water.
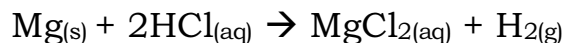
Monobasic a Dibasic acids are acids which have one (mono) or two (di) acidic Hydrogen atoms which can be ionised per molecule. An example of a Monobasic acid would be Hydrochloric Acid as there is only one acidic Hydrogen atom. An example of a Dibasic acid would be Sulphuric Acid as it has two of these Hydrogens. This means it would take 2 moles of a base to neutralise 1 mole of the Dibasic acid.

As my experiment is between strong and weak acids, I will be using Hydrochloric Acid (HCl) and Ethanoic Acid (C_2H_3OOH).

<u>My Experiment - Order</u>

For the order reaction, I will be using strong and weak monobasic acids to test whether the strength of the acid has any effect on the order of the reaction. As I'm using comparing strong and weak acids, I have to make sure that I use two monobasic acids or two dibasic acids. Using one of each will confuse the experiment at I would then be measuring two different things, rather than the one I intend to do. I will therefore be using Hydrochloric Acid (HCl) and Ethanoic Acid (C_2H_3OOH) in this experiment.

As the reaction between a metal and an acid gives off Hydrogen gas, I will be using a gas syringe to measure the rate of this reaction.

$$\text{Metal} + \text{Acid} \rightarrow \text{Salt} + \text{Hydrogen gas}$$

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

$$Mg_{(s)} + 2C_2H_3OOH_{(aq)} \rightarrow Mg(C_2H_3O_2)_{2(aq)} + H_{2(g)}$$

Before beginning my real experiment, I must first do a run of preliminary experiment to test which metal would be most suitable for this reaction, what form the metal should be in and time intervals between measuring the volume of Hydrogen gas produced.

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Calculations

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

First I need to calculate volume of Hydrogen gas that I want to measure. This will help me determine the number of moles of the gas produced.

$$\text{Number of moles} = \frac{\text{Volume of gas}}{24000}$$

By substituting in the numbers that I wanted, I made the equation:

$$\text{Number of moles} = \frac{96}{24000}$$

I then worked out the moles as being 0.004 moles.

With the moles worked out and the ratio of Hydrogen gas to Magnesium being 1:1, I used the mole equation to work out the mass of Magnesium that I needed.

$$\text{Number of moles} = \frac{\text{Mass}}{\text{RMM}}$$

$$0.004 = \frac{\text{Mass}}{24}$$

Therefore, the mass of magnesium that I needed was 0.096g

After working out the moles of the Hydrogen gas, I can also work out the volume of the acid required by looking at the mole ratio (2:1) and then plugging it into another equation:

$$\text{Volume} = \frac{\text{Moles x 1000}}{\text{Concentration}}$$

$$\text{Volume} = \frac{0.008 \text{ x } 1000}{2}$$

The volume of HCl that I need is 4cm^3.

As I need to use the Magnesium in excess so as not to confuse the experiment, I shall be using 0.3g of it instead of 0.096.

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I did the exact same calculation for Ethanoic Acid, only replacing the HCl with the C_2H_3OOH:

$$Mg_{(s)} + 2C_2H_3OOH_{(aq)} \rightarrow Mg(C_2H_3O_2)_{2(aq)} + H_{2(g)}$$

The mass of Magnesium needed was 0.1g, but as I need it in excess, I will be using 0.3g.

The volume of acid needed is 4.16cm^3.

For the Activation Energy experiments I will be using the exact amounts of Magnesium that I need to, 0.1g.

Preliminary Experiments

For my preliminary experiments, I needed to see whether the use of Magnesium of Zinc would be most suitable. Before even undertaking this experiment I knew that I wanted to use Magnesium as it was far higher in the Reactivity table than Zinc, even so, I needed to be sure.

After a test run using 0.1g of Magnesium ribbon and 0.1g of Zinc ribbon, I found that my initial assumptions were correct. The Zinc reacted far too slowly with the acid and the Magnesium gave me just the right reaction.

Metal	Time taken to produce 50cm^3 of H_2(s)
Mg	73
Zn	200

My next experiment was to see whether I should use Magnesium Powder or Magnesium Ribbon. My initial though was to use the powder as it has the highest surface area and would allow the reaction to happen faster...

As I concluded after my first disastrous attempt at using Magnesium Powder in comparison to Magnesium Ribbon, the powder reacted far too quickly for me to monitor effectively and much of the gas escaped before I was able to seal the system.

	Powder	Ribbon
Time taken to produce 70cm^3 of H_2(g)	30	70.6

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Before doing an preliminary experiments, my initial time intervals were 10 seconds apart. After doing the preliminary experiments, I found that it was easier for me to monitor the reaction every 5 seconds apart as I was then able to get more accurate figures. The beginning of the reaction between the Magnesium Ribbon and the acid was very fast and the slowed down towards the end. It is the beginning which is crucial in finding the initial rate of the reaction, so I decided to decrease the time intervals so I was able to get better results.

Activation Energy

For the Activation Energy experiments I had to first find the volume of gas that was the most appropriate to measure. For this, I had to do the experiment with HCl and C_2H_3OOH at a high and low temperature. This was so that I can find a volume which suited both's initial rates of reaction.

HCl –

Time (s)	Volume (cm³) at 20°
0	0
5	6
10	10
15	14
20	18
25	21
30	24
35	25
40	26
45	28
50	29
55	30

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Time (s)	Volume (cm³) at 70°
0	0
5	16
10	21
15	25
20	28
25	30
30	32
35	34
40	37
45	39
50	40

C_2H_3OOH –

Time (s)	Volume (cm³) at 20°
0	0
5	2
10	5
15	7
20	10
25	14
30	20
35	23
40	26
45	30
50	33

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55	36
60	39
65	42
70	45
75	51
80	53
85	56
90	57

Time (s)	Volume (cm³) at 70°
0	0
5	10
10	22
15	35
20	45
25	55
30	60
35	67
40	72
45	76
50	78
55	78

I then drew a graph to illustrate the date above and worked out a suitable amount of Hydrogen gas to measure:
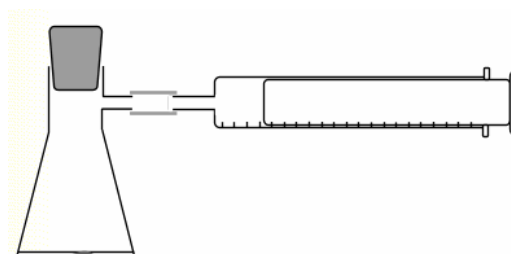
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After drawing the graphs and finding suitable volumes to measure for each, I had to decide which temperature I wanted it to be measured at. My initial though was every 20 degrees: 20 °, 40 °, 60 ° and 80 °. After a test run for each temperature, I found that these temperatures were perfect to use. My one problem was, however, the reaction occurring too fast for me to monitor.

I then decided to dilute the 2M of acid into 1M by adding 4cm^3 of water and halving the volume. The addition of water without the removal of half was doubling the volume as well as diluting the acid. If I were to use the volume without halving, I'd still be using 1M of 8cm^3 of acid. That is why I halved the volume.

<u>Apparatus</u>

- Small Measuring Cylinder

- Water Bath

- Hydrochloric and Ethanoic Acid – correct amounts

- Magnesium Ribbon – Correct amounts

- Kettle

- Gas-Syringe

- Boiling Tubes

- Stop-Clock

- Distilled Water

- Clamp Stand

- Test-Tube Rack

<u>Methods</u>

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Order –

1) First gather all of the equipment above and make sure the correct amounts of each substances are ready.

2) Ensure that the syringe is clamped securely and that it is exactly horizontal. The values on the syringe need to be shown clearly so that you can measure the amount of gas produced when the experiment is up and running.

3) Spiral the Magnesium Ribbon pieces around a pencil and do the same for the other pieces.

4) Add the correct amount of acid to the boiling tube and place in a water bath under the syringe, ready for the experiment to begin.

5) Drop the Magnesium Ribbon carefully into the Acid and start the stop-clock as soon as the Ribbon hits the acid.

6) Monitor the experiment every 5 seconds and record the volume at that time in the results table.

7) Repeat the experiment another two times so that you are able to find the average volume for the three experiments.

8) Do the same experiment for the other acid – ensuring that you repeat it until you have 3 values for each.

Results Table for Order

HCl –

Time (s)	Volume (cm³)				Vf-Vt
	1	2	3	Average	
0	0	0	0	0	93
5	41	35	40	38	55
10	59	55	53	56	37
15	65	63	63	64	29
20	71	72	71	71	22
25	74	75	77	76	17

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30		77	80	79	14
35		79	85	82	11
40		80	87	85	8
45			90	90	3
50			91	91	2
55			93	93	0

C_2H_3OOH –

	Volume (cm³)				
Time (s)	1	2	3	Average	Vf-Vt
0	0	0	0	0	81
5	2	2	2	2	79
10	4	4	4	4	77
15	8	5	7	7	74
20	10	10	10	10	71
25	17	15	17	16	65
30	22	20	20	21	60
35	30	27	26	27	54
40	37	33	30	33	48
45	47	39	39	42	39
50	60	45	47	50	31
55	70	54	59	60	21
60	78	58	62	67	14
65	80	60	69	70	11
70	82	62	75	73	8
75	85	64	79	76	5
80	87	65	83	79	2

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85	90	65	85	80	1
90	92	65	85	81	0

Methods

Activation energy –

1) Like for the order reaction, place the boiling tube sin separate water baths.

2) Heat up the kettle and pour into 3 of the water baths – normal tap water in the last.

3) Put a thermometer in each acid and keep an eye on the temperature of each so as not to let it get passed the one you need.

4) Start with the hottest acid and record how long it takes the syringe to get to the set volume, 15cm³ for HCl and 30 cm³ for C_2H_3OOH. Repeat the with the same temperature 2 more times until you have 3 values for it.

5) Work through the acids as they reach the correct temperature, ensuring you add more hot/cold water depending on the current temperature. Like the above, repeat the experiment with each temperature until you have 3 values for each.

6) Do the same for the other acid, ensuring that each temperature has 3 values.

Results Table for Activation Energy

HCl –

Temp (°C)	Temp (K)	1/T (K)	Time (s)				1/t (s-1)	ln(rate)
			1	2	3	Avg		
20	293	3.41×10^{-3}	11	11	11	11	0.09	-2.41
42	315	3.17×10^{-3}	9	7	9	9	0.11	-2.21
60	333	3.00×10^{-3}	7	8	8	8	0.125	-2.08

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80	353	2.83 x 10⁻³	5	5	6	5	0.20	-1.61

C$_2$H$_3$OOH –

Temp (°C)	Temp (K)	1/T (K)	Time (s)				1/t (s⁻¹)	ln(rate)
			1	2	3	Avg		
20	293	3.41 x 10⁻³	77	75	75	75	0.013	-4.34
39	312	3.21 x 10⁻³	25	26	25	25	0.04	-3.22
60	333	3.00 x 10⁻³	10	10	11	10	0.10	-2.30
77	350	2.86 x 10⁻³	6	6	7	6	0.167	-1.79

Safety

As I am using acids which are known to be corrosive and irritating to the skin, some safety precautions need to be taken:

- Goggles must be worn at all times when working on an experiment. This is due to the irritating nature of the strong acids used, as well as the vigorous reaction that may occur between the Magnesium ribbon and HCl acid.

- Like the goggles above, lab-coats must be worn at all times to ensure that no spillage comes into contact with the skin or with clothes.

- Gloves must be worn when handling the acids to stop them from coming into contact with the skin.

- One must take when handling the hot water from the kettle as glassware gets very hot very quickly. To avoid any burns, everything must be handled with care.

- Any acid spillages must be cleaned up quickly to stop anyone from accidentally leaning on it and injuring themselves.

- As Hydrogen is the being produced, ensure that no Bunsen Burners are lit within the vicinity. Hydrogen is a very flammable gas and lighting a flame in a closed area like the lab is very dangerous.

Fair Test

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Throughout the experiment there are a number of variables that will affect the outcome of the experiment. To keep these variables at a minimum, the following must be done:

- The amount of reactants used MUST be a constant. If one piece of Magnesium is 0.11g, then the other pieces need to be 0.11g. The amount of acid used must be the same – if 4cm^3 of acid is used for one experiment, it must remain 4cm^3 for the rest of the duration.

- The way the Magnesium is twisted must be the same for each piece used. If you use a pencil for one of the pieces, you must use the same pencil for the rest of them.

- There is a thin layer of Magnesium Oxide coating the Magnesium Ribbon. This will interfere with the experiment so you need to sand down the Magnesium to get rid of this layer. However, if you haven't sanded down the Magnesium for the previous experiments, then you mustn't do it for the proceeding ones. You can, however, repeat the previous experiments and sand the Magnesium for those.

- As the reactions are exothermic, the temperature is sure to rise. As stated at the beginning of this coursework, an increase in temperature increases the rate at which the reaction happens. As the temperature isn't a variable for the Order reaction, the boiling tubes must be put into a water bath at room temperature (20 º). For the activation energy experiments, as the temperature is a variable, an outside increase in the temperature will throw the experiment off balance. That is why the boiling tubes are kept in the water bath with the correct temperature, rather than heated with a flame.

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Analysis of Order

The first graph I drew was for the reaction between 2M of HCl and the Magnesium Ribbon. After drawing the successive half-lives, I deduced the order as being the First order. My values were slightly off, and I put that down to slight errors in the experiment, as the half-lives I got were 16 seconds, 14 seconds and 12 seconds. The margins between the 3 half-lives aren't that great, however, so one can still assume that the order is the first.

The anomalies for the HCl graph weren't really very there as they all pretty much fitted with the line of best fit.

The second graph was for the reaction between 2M of C_2H_3OOH and the Magnesium ribbon. After drawing the graph, I saw that there were many anomalies within the results and I had to draw the line I felt best fitted the graph. With this line of best fit, I was able to deduce that the order of this reaction was also the first order. The successive half-lives, like for the HCl graph, weren't all equal, but the margins between the first 2 is close enough. The half-lives were 22 seconds, 20 seconds and 16 seconds. As you can see, the first 2 were as close together as for HCl, so one can just assume that the reaction was the first order.

The anomalies for the C_2H_3OOH graph were a lot more apparent. I am unable to explain why the results came how they did as this particular experiment was repeat another 2 times with the same results – I can only deduce that there was some experimental equipment errors along the line which we may have overlooked. The line of best fit that I drew was the line I assumed the graph would follow, had the results been correct.

After drawing the graphs and deducing the orders from what the graphs illustrate, I believe the order of the reaction with respects to the acids were the First order. This is due to both of the graphs having similar enough successive half-lives.

Conclusion of Order

As I have found out that both of the reactions are in the first order, I can say that the strength of the acid has no effect of the order of the reaction. While the half-lives aren't equal, they are similar enough to not be mistaken as the second order – this is why I can assume that the strength of the acid has no effect on the order of the reaction.

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The rate determining step of a reaction is the slowest step that occurs in a reaction. As the reaction can't progress until the slowest step has been passed, the rate of the reaction depends on this step.

In the reaction between an acid and a metal, many steps need to be taken:

- A positive Hydrogen ion will need to be disassociated from the acid. This ion will then react with the water to produce an Oxonium ion.

- The metal needs to lose an electron and this is done when the Oxonium ion reacts with the it.

- The electron will join with the positive Hydrogen ion and form Hydrogen gas.

- The Hydrogen molecule will then move from the active site.

$$H_3O^+ + Mg \rightarrow Mg^+ + H_2O + H$$

The slowest step from the above would be the reaction between the Oxonium ion and the Magnesium metal as it requires the most energy. This is because extracting electrons from the Magnesium metal is difficult. After the ion has enough energy to attract the electron, it will then ionise the metal. The other steps can't be done without this step being completed and as this step required the most amount of energy and is the most difficult one, it is the rate determining step.

Analysis of Activation Energy

After drawing the activation energy graphs, finding the gradient of them and plugging the values into the Arrhenious equation, I was able to gain the Ea values for HCl and C_2H_3OOH.

HCl = 67.76 kJ K^{-1}

C_2H_3OOH = 326.46 kJ K^{-1}

The only anomaly present in the two of the graphs was for the HCl one. An anomaly occurred at 80 °, but it had no effect on the line of best fit.

Conclusion of Activation Energy

From the results above, I can conclude that the strength of the acid is inversely proportional to the activation energy. Ethanoic Acid (C_2H_3OOH), a weak acid, has a very high activation energy; much higher than Hydrochloric Acid (HCl), a strong acid with a small value of 67.67 kJ K^{-1}.

The results show that weak acids take longer to react as they have a much higher activation energy to reach. Strong acids, such as HCl, have a much

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smaller activation energy which can explain why they react so rapidly when in high concentrations.

As stated previously, the strength of an acid depends on its ability to disassociate its acidic Hydrogen protons in water. The production of the Oxonium ion is one that can help to measure the strength of an acid.

As the reaction to produce the Oxonium ion is a reversible one and Ethanoic acid only partially ionises in water, it is safe to assume that once the Oxonium ion has been produced with the Ethanoic acid, it reverts back into the acid as it has become unstable due to its much slow ionisation.

$$C_2H_3COOH_{(aq)} + H_2O_{(l)} \rightarrow C_2H_3COO^+_{(aq)} + H_3O^+_{(aq)}$$
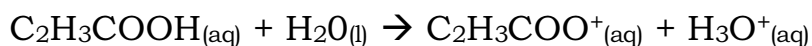
<u>Evaluation</u>

The experiment I chose to conduct was to investigate the effects the strength of the acid had on the order of a reaction and its activation energy. As the equipment I was using throughout this experiment was far from expertly accurate, there were many errors within the experiment that could account for some of the anomalies I found in my order experiments. Although the equipment used wasn't terrible, the margins of errors could have accumulated over the course of the experiment. Human errors, as well as experimental errors, would also have been made.

Firstly, the measuring cylinders used to measure out the 4cm³ of Hydrochloric and Ethanoic Acid would only measure to nearest 0.05cm³ and this can account to a ±1.25% error. Although this isn't a huge error, it is significant due to the small volumes I was using.

The gas-syringes only measured to the nearest cm³ and this left an error of ±0.5 cm³. As I was measuring how long it took for 15 cm³ of Hydrogen gas to be produced, that 0.5cm³ error would amount to a ±3.33% error. This is hugely significant due to the small volume I was measuring.

The balance I was using to weigh the Magnesium Ribbon only measure to nearest 0.01g. For the order experiments, accuracy wasn't an issue as I was using the Magnesium in excess. However, for the activation energy experiments I was using exact values and needed 0.1g of Magnesium Ribbon. There was a potential error of ±5%.

The thermometers we were using would only measure to the nearest degree; this leaves an chance of the value being ±0.5°. This is a rather small value when taking into consideration the large temperatures we were using, so this error is negligible. One concern of mine was how quickly the temperatures moved from being the correct one to just under. The start of the experiment saw the temperature as being 77° and the end of the experiment saw it as

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being 74 °. This is a very large change and could account for some of the errors made.

Another error I found was not sealing the system fast enough. This was most apparent when using the highest temperatures and during the preliminary experiments with the Magnesium Powder. Not getting the bung on quickly enough meant that some of the gas escaped before being sealed in the syringe and so the initial rate of the reaction would be skewed. As the reaction in higher temperatures happened rather vigorously, once the first few seconds have passed, the reaction may have reached its peak already. Not sealing the tube quickly enough would mean that an accurate reading of the results wouldn't have been possible to obtain.

If I were to conduct this experiment again, there would be some errors that could possibly be eliminated. Getting another person to start the stop-clock whilst I drop the Magnesium into the reaction and seal it with a bung would be the first correction. This would severely reduce the amount of gas which had escaped from the system in my previous experiment. The Oxide layer coating the Magnesium, although previously sanded, would have inevitably built up during the course of the experiment. Sealing the pieces in an air-tight container would reduce any reaction from occurring before the pieces are used. Using a burette to measure out the acid volumes rather than a measuring cylinder would definitely make the quantities more accurate and the results more reliable. Using water baths like I had previously to keep the temperature constant would be used again to stop any outside temperatures from affecting the rate of the reaction.

I would also change my experiment slightly and use Mono and Dibasic acids instead. This way, I could potentially see whether two strong acids with a different number of acidic Hydrogens react differently and affect the order and activation energy significantly. Using Sulphuric acid and Hydrochloric acid would be my choice.