Chemistry Knowledge Organiser

Topic 5: Energetics & Rates of Reaction

Key Terms	Definitions
Reaction Profile	A graph which shows the energies of the products and reactants in a chemical reaction
Exothermic	A reaction that gives out heat to the surroundings
Endothermic	A reaction that takes heat in from the surroundings

Energy in Reactions

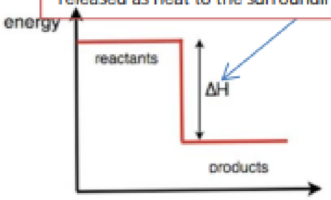
Energy is conserved in chemical reactions. The amount of energy in the universe at the end of a chemical reaction is the same as before the reaction takes place. In a chemical reaction, bond breaking and bond making occur. To break a chemical bond you need to overcome the force of attraction in the bond, this process requires energy therefore it is endothermic. The process of bond formation is exothermic, energy is released when bonds form. In a chemical reaction the difference between the energy required to break the bonds and the energy gained from making the bonds will decide whether a reaction is exothermic or endothermic.
Chemical reactions can therefore be divided into exothermic and endothermic chemical reactions.

	What happens?	Why?	Example
Exothermic	Heat energy is transferred to the surroundings.	The energy required to break chemical bonds is less than the energy gained from making chemical bonds. Therefore the excess is given off as heat to the surroundings.	Combustion reaction, reactions used in hand warmers
Endothermic	Heat energy is taken in from the surroundings	The energy required to break chemical bonds is more than the energy gained from making chemical bonds. Therefore heat is taken in from the surroundings.	The reaction of citric acid and sodium hydrogencarbonate, the reactions used in ice packs

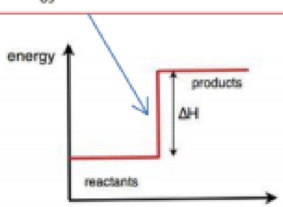
Reaction Profiles

Reminder from topic 15: Chemical reactions can occur only when reacting particles collide with each other and with sufficient energy. The minimum amount of energy that particles must have to react is called the activation energy. Reaction profiles can be used to show the relative energies of reactants and products, the activation energy and the overall energy change of a reaction.

This is the reaction profile of an exothermic reaction, the energy of the products is lower than that of the reactants. The difference in energy is released as heat to the surroundings.



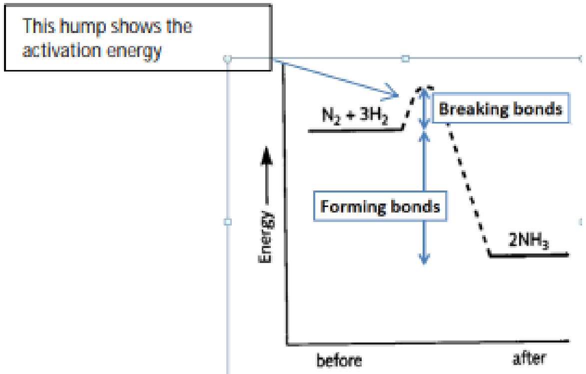
This is the reaction profile of an endothermic reaction, the energy of the products is higher than that of the reactants. The difference in energy is taken in from the surroundings.



Reaction Profiles- In more detail

The profile below shows the reaction which makes ammonia from nitrogen and hydrogen. The equation is given below:

$$N_2 + 3H_2 \rightarrow 2\ NH_3$$



There are some key features to highlight on this graph ,firstly the curved section represents the activation energy for this reaction, this hump shows how much energy is required to break the bonds in the reactants. To overcome the activation energy we often need to heat our reactants. The products are lower in energy than the reactants, this means it is an exothermic reaction. As the excess energy is given out to the surroundings. as heat energy.

Calculating bond energies -higher tier.

The difference between the sum of the energy needed to break bonds in the reactants and the sum of the energy released when bonds in the products are formed is the overall energy change of the reaction.
For example consider the reaction:

$$N_2 + 3H_2 \rightarrow 2\ NH_3$$

To work out the overall energy change you will need to subtract, the energy gained from forming the bonds in ammonia, from the energy required to break the nitrogen and hydrogen bonds. This will give you the overall energy change, if the value is negative then the reaction is exothermic, if the value is positive the reaction is endothermic

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Topic 5: Energetics & Rates of Reaction

Bond Energies continued- Higher Tier

You can calculate the energy change in a reaction from bond energies given to you in a question. For example consider the reaction below:

2 H – O – O – H ⟶ 2 H – O – H + O = O

This shows that hydrogen peroxide breaks down to make water and oxygen. We can use bond energies to work out the energy change in the reaction.

Bond	Bond energy in kJ per mole
H O	464
O O	146
O = O	498

The energy required to break the reactant bonds is:

2(H – O)(O – H)

2 x464 (for the O-H bonds) = 928 + 146 (for the O-O bond)=1074 however as there is a 2 in the equation this number needs to be doubled.

2 x 1074 = 2148 klj/mol

The energy gained from making the product bonds is:

2 H–O–H O=O

2x464= 928 but there is a 2 in the equation so this doubled to 1856 and we also need to add the 498 for the double bond in O_2

1856+498=2354 kj/mol

Therefore we do energy required to break reactant bonds- energy gained from making product bonds:

2184-2354=-170kj/mole

If the value is negative then the reaction is exothermic
If the value is positive the reaction is endothermic.

Key Terms	Definitions
Equilibrium	A reaction that is reversible
Le Chatelier's principle	A principle which states, "If a system is at equilibrium and a change is made to any of the conditions, then the system responds to **counteract the change"**
Dynamic Equilibrium	An equilibrium where the forward and backward reactions are happening at the same rate

Equilibrium

Some chemical reactions are reversible, this means they can happen in both the forward and reverse directions. The symbol we use to represent an equilibrium reaction is shown in the equation below:

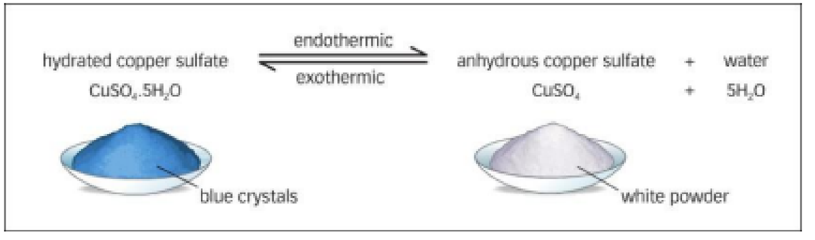
$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

In a dynamic equilibrium reaction, the forward and reverse reactions are happening at the same rate.

A dynamic equilibrium has to occur in a closed system, where no reactants and products are allowed to escape.

If the equilibrium lies to the left, it means that there is a greater concentration of reactants than products
If the equilibrium lies to the right it means there is a greater concentration of products than reactants.

Most equilibrium reactions are endothermic in one direction and exothermic in another direction. A good example is the hydration and dehydration of copper sulphate. It is exothermic when water is added to the copper sulphate, it is endothermic when water is removed.



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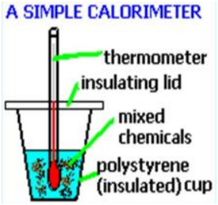
Required Practical
Chemistry – Temperature changes

Objective: Investigate the variables that affect temperature change in chemical reactions eg acid plus alkali.

In this practical you will:

- react sodium hydroxide solution with hydrochloric acid.
- measure the temperature changes during the reaction.
- plot a graph of your results and record the temperature change.

Method



1. Measure 30cm^3 dilute hydrochloric acid and put it into the polystyrene cup.
2. Stand the cup inside the beaker. This will make it more stable.
3. Use the thermometer to measure the temperature of the acid. Record your result in a table like this.
4. Measure 5cm^3 sodium hydroxide solution.
5. Pour the sodium hydroxide into the polystyrene cup. Fit the lid and gently stir the solution with the thermometer through the hole.
6. Look carefully at the temperature rise on the thermometer.
7. When the reading on the thermometer stops changing, record the highest temperature reached in the table.
8. Repeat steps 4–7 to add further 5cm^3 amounts of sodium hydroxide to the cup each time, recording your temperature reading in the results table.
9. Repeat until a maximum of 40cm^3 of sodium hydroxide has been added.
10. Wash out all the equipment and repeat the experiment for your second trial.

Total volume of sodium hydroxide added in cm^3	Maximum temperature in °C		
	First trial	Second trial	Mean
0			
5			
10			

Apparatus

- dilute hydrochloric acid
- dilute sodium hydroxide solution
- an expanded polystyrene cup and lid
- 250 cm^3 beaker
- 10 cm^3 measuring cylinder
- 50 cm^3 measuring cylinder
- a thermometer

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Required Practical
Chemistry – Temperature changes

Objective: Investigate the variables that affect temperature change in chemical reactions eg acid plus alkali.

Total volume of sodium hydroxide added in cm^3	Maximum temperature in °C		
	First trial	Second trial	Mean
0	20.0	21.0	
5	24.0	24.6	
10	26.8	27.6	
15	28.6	29.6	
20	30.8	31.3	
25	31.8	32.8	
30	32.0	32.6	
35	31.6	31.8	
40	30.6	31.0	

Analysis and conclusion

a. Calculate the mean maximum temperature reached for each volume of sodium hydroxide. Record these means in your table.

b. Plot a graph from your results and draw two straight lines of best fit.

c. From the graph read off the maximum temperature change.

d. This is an example of an exothermic reaction, when heat is given out. Can you explain why the results you recorded show that the temperature started to fall after a certain volume of sodium hydroxide had been added?

e. Can you explain what is happening to the chemical bonds in the reactants and the products when an exothermic reaction is taking place?

Health and safety

- Eye protection
- sulfuric acid: IRRITANT
- hydrochloric acid: IRRITANT

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Topic 5: Energetics & Rates of Reaction

Changing Conditions-Le **Chatelier's** principle- Higher Tier

The Haber process is a good example to explain Le **Chatelier's** principle, the equation for the Haber process is shown below. The reaction is carried out in the gaseous state. Remember this is one of many reactions but the principles always stay the same.

Endothermic in this direction $N_2 + 3H_2 \rightleftharpoons 2NH_3$ Exothermic in this direction

Condition Change	Effect
Increase the temperature	Shifts the equilibrium to the left as this is the endothermic direction. The amount of reacrtants increases.
Decrease the temperature	Shifts the equilibrium to the right as this is the exothermic direction. The amopunt of product increases
Increase the concentration of reactants	Equilibrium shifts to the right to make more product, to reach equilibrium again
Increase the concentration of products	Equilibrium shifts to the left to reach equilibrium again
Increase the pressure in the gas	Equilibrium shifts to the right, where there are fewer molecules of gas, this will decrease the pressure.
Decrease the pressure in the gas	Shifts the equilibrium to the left as there are more gas molecules on that side of the equation.

Key Terms	Definitions
Equilibrium	A reaction that is reversible
Le Chatelier's principle	A principle which states, "If a system is at equilibrium and a change is made to any of the conditions, then the system responds to **counteract the change"**
Dynamic Equilibrium	An equilibrium where the forward and backward reactions are happening at the same rate

Equilibrium- Changing Conditions-Higher tier

The amounts of all the reactants and products at equilibrium depend on the conditions of the reaction. For example if we change things like temperature, concentration of a reactant or product and pressure in gases.

The French scientist Le Chatelier devised a principle to explain how equilibrium reactions, respond to a change in conditions, it states that:

"If a system is at equilibrium and a change is made to any of the conditions, then the system responds to counteract the change"

For example if the temperature is raised the equilibrium will shift to try to cool the surroundings down.

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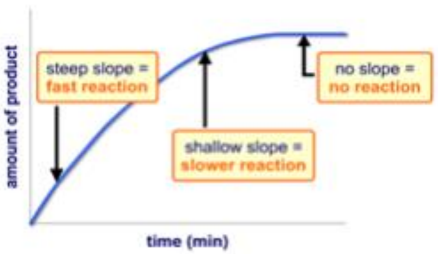
Topic 5: Energetics & Rates of Reaction

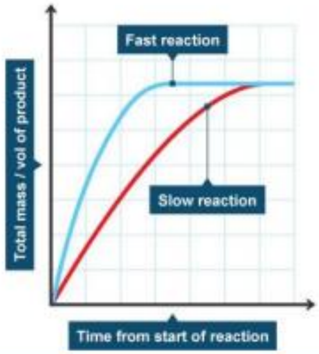
Key Terms	Definitions
Activation Energy	The minimum energy required for a chemical reaction to take place
Collision Theory	The theory that states for a chemical reaction to happen, particles must collide with sufficient energy
Gradient	The measurement of how steep a line is on a graph
Frequency	The amount of times something happens in one second
Concentration	The number of particles in a given volume

Interpreting Rate of Reaction Graphs

The results from rate of reaction experiments can be plotted on a line graph. For example how the mass changes against time or how much gas is made against time. Different lines can be plotted for different conditions, the steeper the gradient, the faster the reaction.

It is important to remember that the graphs flatten off (plateau) at the same point as the same amount of reactant is being used.





Factors which affect Rate of Reaction

Being able to slow down and speed up chemical reactions is important in everyday life and in industry. We can change the rate of a reaction by:

- Changing temperature
- Changing pressure
- Changing the concentration of a solution
- Changing the surface area
- Adding a catalyst

Collision Theory

Collision Theory: reactions occur when particles collide with a certain amount of energy.

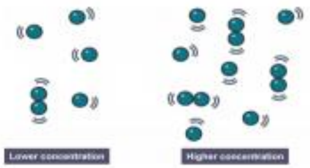
The minimum amount of energy needed for the particles to react is called the activation energy, which is different for each reaction.

The rate of a reaction depends on two things:

- the frequency of collisions between particles. The more often particles collide, the more likely they are to react.
- the energy with which particles collide. If particles collide with less energy than the activation energy, they will not react.

Collision Theory- in more detail Concentration

If the concentration of a solution is increased then there are more particles in a given volume, therefore collisions are more frequent and the chemical reaction is faster. Concentration is directly proportional to rate of reaction (if you double the concentration you double the rate



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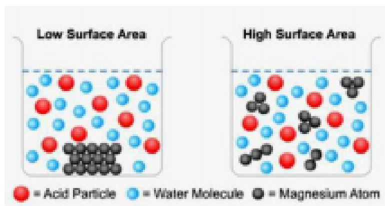
Topic 5: Energetics & Rates of Reaction

Collision Theory in more detail Temperature

When you increase the temperature of something the particles will move around faster, this increases the frequency of the collisions. As well as that, as the particles are moving faster the particles collide with more energy making it more likely that collisions exceed the activation energy.

Collision Theory in more detail Surface Area

When you increase the surface area of a solid (you cannot increase the surface area of a liquid or gas). You increase the number of particles that are available for collision, therefore increasing the frequency of collisions therefore increase the rate of reaction.



Collision Theory- in more detail gas pressure

If the reaction is carried out in the gaseous state, then increasing the pressure will increase the rate of reaction. If there are more particles in a given volume of gas, then collisions will be more frequent and therefore the reaction will be faster.

(a) Low pressure (b) High pressure

Enzymes

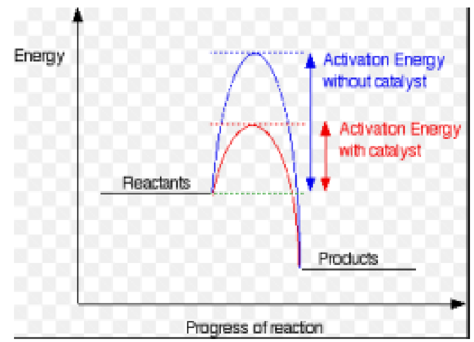
Enzymes are biological catalysts, they speed up chemical reactions in biological systems for example in digestion in animals. Unlike catalysts enzymes have an optimum temperature where they work best, this is usually around 37

Key Terms	Definitions
Enzymes	A biological catalyst
Reaction Profile	A graph which show the energies of the reactants and products at different stages of the chemical reaction

Collision Theory in more detail Catalysts

A catalyst is a substance which speeds up a chemical reaction without being used up. It speeds up a reaction because it lowers the activation energy by providing an alternative pathway and this means that there are more successful collisions and a faster reaction.

The effect of a catalyst is shown on the reaction profile below:



Catalysts are not included in a chemical equation as they are not used up in a chemical reaction.

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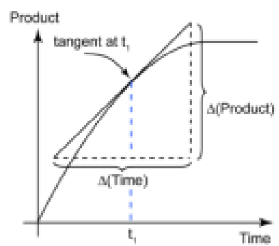
Topic 5: Energetics & Rates of Reaction

Rate of Reaction

The rate of reaction is the speed at which a chemical reaction is happening. This can vary hugely from reaction to reaction.

The rate of reaction can be calculated either by measuring the quantity of reactant used or the quantity of product made in a certain length of time. The quantity can either be a volume measured in cm^3 or a mass measure in grams (g).

Measuring Rate of Reaction-Higher Tier

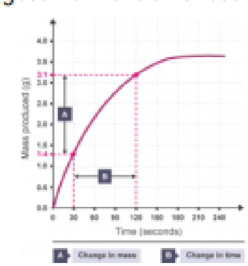


The gradient of a volume or mass/time graph will give you the rate of reaction at a given point. However when the line is a curve you need to draw a tangent to measure the gradient. To draw a tangent follow the following steps

1. Line you ruler up across your graph, so that it touches the line on the point that you want to find out the gradient
2. Adjust the ruler until the space between the ruler and the curve is equal on both sides
3. Draw the line and pick two easy pints that will allow you to calculate the gradient of the line.

Calculating the Mean Rate of Reaction -Higher Tier

To calculate the mean rate of reaction from a graph you need to pick two y values on the graph and two x values, subtract the largest from the smallest and the divide the value on the y axis by the valued on the x axis.



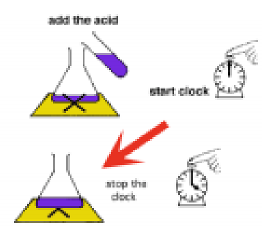
Key Terms	Definitions
Rate of Reaction	The rate at which reactants are being turned into products
Reactant	What is used in a chemical reaction
Product	What is made in a chemical reaction
Catalyst	A substance which speeds up a chemical reaction without being used up
Tangent	A straight line that touches a curve at a point

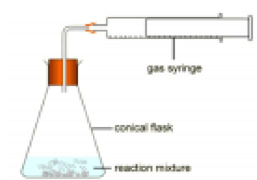
Equation	Meanings of terms in equation
Rate of Reaction= $\frac{Reactant\ used}{time}$	*Reactant used can either r be measured in grams or cm^3*
Rate of Reaction= $\frac{Product\ \ Made}{time}$	*Reactant used can either r be measured in grams or cm^3*

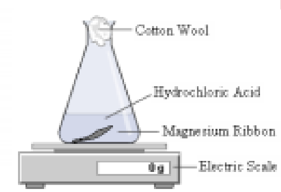
Measuring the Rate of Reaction

There are several experiments that can be used to measure the rate of a chemical reaction.

1. Measuring the mass lost in a chemical reaction (marble chips and acid is a good example)
2. Measuring the volume of gas produced (decomposition of hydrogen peroxide is a good example)
3. Time taken to make an X disappear (sodium thiosulphate and acid is a good example)







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Required Practical
Chemistry – Concentration and rate of reaction.
Activity 1

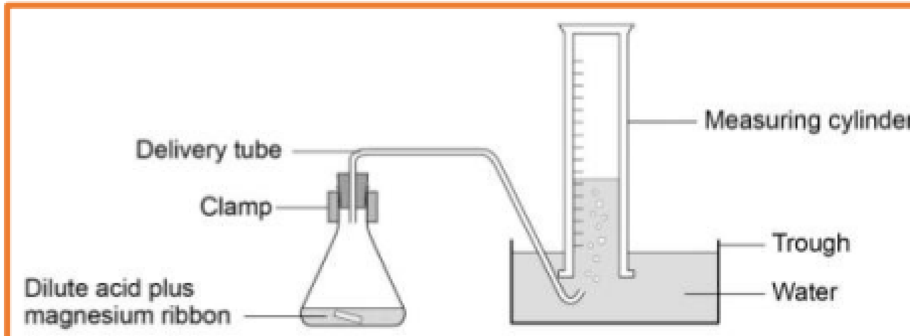
Objective: Investigate how changes in concentration affect the rates of reactions by measuring the volume of a gas produced.

In this practical you will:
- react magnesium ribbon with different concentrations of hydrochloric acid.
- measure the volume of gas produced for each concentration.
- use your results to work out how the rate of reaction is affected by the concentration of the acid.

Method
1. Measure 50cm³ of 1.0m hydrochloric acid using one of the measuring cylinders. Pour the acid into the 100cm3 conical flask.
2. Fit the bung and delivery tube to the top of the flask.
3. Half fill the trough or bowl with water.
4. Fill the other measuring cylinder with water. Make sure it stays filled with water when you invert it into the water trough and that the delivery tube is positioned correctly.
5. Add a single 3cm strip of magnesium ribbon to the flask, put the bung back into the flask as quickly as you can, and start the stop clock.
6. Record the volume of hydrogen gas given off at suitable intervals (eg 10 seconds) in a table.
7. Repeat steps 1–6 using 1.5m hydrochloric acid.

Time in seconds	Volume of gas produced cm^3	
	1.0 M	1.5 M
10		
20		
30		
40		
50		
60		
70		
80		
90		
100		

Analysis, conclusion and evaluation
a. Plot a graph of your results.
b. Draw a line of best fit. The results should generate a curve not a straight line.
c. Plot the curve for both 1.0M and 1.5M hydrochloric acid on the same graph.
d. Use your graph to compare the rates of reaction with different concentrations of hydrochloric acid with magnesium.
e. Use kinetic theory to explain your findings.



Required Practical
Chemistry – Concentration and rate of reaction.
Activity 2

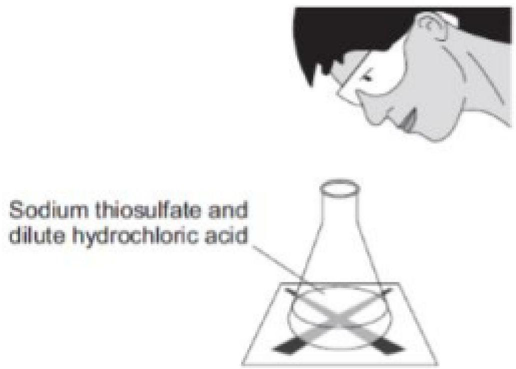
Objective: Investigate how changes in concentration affect the rates of reactions by monitoring a change in colour or turbidity.

In this practical you will:

- react different concentrations of sodium thiosulfate with hydrochloric acid.
- use a stop clock to time how long it takes for the mixture to become cloudy for each concentration.
- use your results to work out how the rate of reaction changes as the concentration of the sodium thiosulfate changes.

Method

1. Measure 10cm3 sodium thiosulfate solution and put it into the conical flask.
2. Measure 40cm^3 of water. Add the water to the conical flask.
3. This dilutes the sodium thiosulfate solution to a concentration of 8 g/dm^3.
4. Put the conical flask on the black cross.
5. Measure 10cm^3 of dilute hydrochloric acid.
6. Put this acid into the flask. At the same time swirl the flask gently and start the stop clock.
7. Look down through the top of the flask. Stop the clock when you can no longer see the cross.
8. Record the time it takes for the cross to disappear in the table below. Record the time in seconds.
9. Repeat steps 1–7 changing the concentration of sodium thiosulphate each time as below
    1. 20cm^3 sodium thiosulfate + 30cm^3 water (concentration 16g/dm^3)
    2. 30cm^3 sodium thiosulfate + 20cm^3 water (concentration 24g/dm^3)
    3. 40cm^3 sodium thiosulfate + 10cm^3 water (concentration 32g/dm^3)
    4. 50cm^3 sodium thiosulfate + no water (concentration 40g/dm^3)



Concentration of sodium thiosulfate in g/dm^3	Time taken for cross to disappear in seconds			
	First trial	Second trial	Third trial	Mean
8				
16				
24				
32				
40				

Analysis, conclusion and evaluation

a. Share results with two other groups. Record these results in the second and third blank columns of your table.
b. Calculate the mean time for each of the sodium thiosulfate concentrations. Leave out anomalous values from your calculations.
c. Plot a graph and draw a smooth curved line of best fit.
d. Describe the relationship between the independent variable and the dependent variable? What were your control variables?
e. Compare your results with those of others in the class. Is there evidence that this investigation is reproducible?
f. Evaluate the two methods that you have used to investigate the effect of concentration on rate of reaction.