Chemical Kinetics

Contents

Kinetics: rates of reaction

The word *kinetics* comes from the Greek word meaning movement (kinein "to move") and we see it in modern words like cinema. In science, it is the branch of physical chemistry that deals with the speed of reactions and the factors that affect how fast reactions occur. We talk about the rate of reaction, the change of something with time which in the case of a chemical reaction is the concentration of reactants or products. Not all reactions are the same and everyday experience tells us that reactions occur at different rates.

Can you think of some examples of every day reactions that occur at different rates?

- Explosions
- Frying an egg – it doesn't fry on its own!
- The sound of crickets on a summer evening.

Kinetics and Thermodynamics.

The kinetics of a reaction are independent of the thermodynamic feasibility of the reaction. Thermodynamics tells us how far the reaction will go, kinetics tells us how fast it will happen. For example, you are made up of a lot of carbon and hydrogen and you move around in the presence of enough oxygen to have complete combustion. From a thermodynamic point of view, there are all the ingredients for a nice exothermic reaction; however, you do not burst into flames spontaneously (well not usually…). This is because of the kinetics of the reactions involved. Studying kinetics tells us why.

The study of chemical kinetics is very important in understanding enzymes and chemical reactions on an industrial scale where it is important to keep reactions under control.

How can we study the kinetics of a reaction in a laboratory?

➢ With your partners discuss how you could study the speed of a reaction.

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Typically, we set up reactions and analyse them at time intervals:

- We can watch the appearance of products or the disappearance of reactants for example through a change in some property of the reaction that we can measure….
  - volume - gas produced
  - colour
 - conductivity
  - formation of a precipitate

From our observations we draw graphs of concentration against time to calculate reaction rates under particular conditions of reaction. The ability to draw and interpret graphs is very important in the study of chemical kinetics.

Exercise in class: the LEGO simulation – introducing the graphical method

Lego kinetics instructions

At the start, the bag contains 20 molecules, 10 each of two different colours (each represents a chemical species).

The molecules can be combined in several ways:
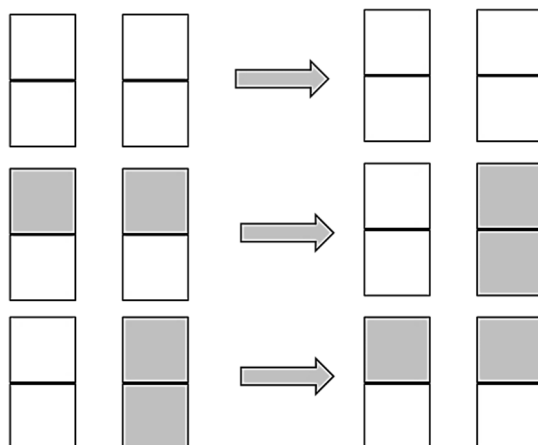
Round one = corresponds to the first time period.

- Extract two molecules.
- Make the product molecules from this combination (see the diagram above). IF THE MOLECULES ARE BOTH OF MIXED TYPE (AB), THE REACTION MUST GO BACKWARDS! (Example 2 above)
- Put them back in the bag.
- Count the TOTAL NUMBER OF PRODUCT MOLECULES PRESENT IN THE BAG before proceeding to the next step.

Round two – twenty five

- Repeat as above – ALWAYS COUNT THE TOTAL NUMBER OF PRODUCT MOLECULES IN THE BAG BEFORE PROCEEDING

For each go, record the total number of product molecules in the bag before proceeding to the next turn.

Turn	Number of product molecules	Turn	Number of product molecules	Turn	Number of product molecules
1		11		21	
2		12		22	
3		13		23	
4		14		24	
5		15		25	
6		16			
7		17			
8		18			
9		19			
10		20			

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Reflection

- Describe the rules of the game.

    o   What type of chemical reaction are we simulating in class?

- Write a balanced equation for this reaction.

    o   Find a real-life example of the same type of reaction.

- Describe analytically the situation at the start of the game.

- What are you simulating when you pick two "molecules" from the bag?

- Did you get the same as the other groups in the class?

- Why/ Why not?

- Which branch of mathematics is at the heart of this simulation?

Results

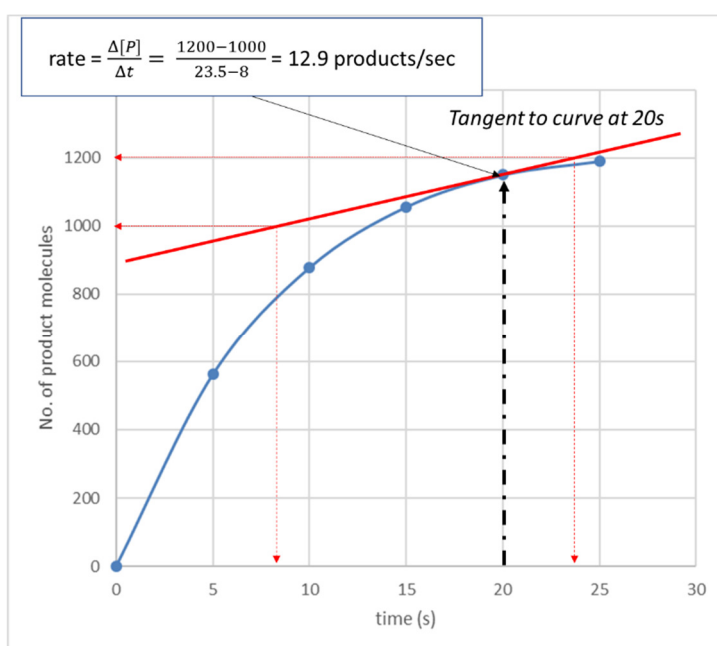The results of a large number of groups are shown below.

Time (s)	Number of product molecules
0	0
5	562
10	878
15	1056
20	1150
25	1190

- Draw a graph of this data using suitably labelled axes. Describe the graph.
- Using the graph. Measure the initial rate of the reaction (time = 0) by *the tangent method*. How does this compare to the rate at 10 minutes and 20 minutes?
- In this experiment we measured the production of product molecules. What would the curve for one of the reactant molecules look like?
- Using the rules that were given at the beginning, does this reaction end with the total disappearance of the reactants?
- How could you modify the game to simulate:
    1. a complete reaction.
    2. Other factors involved in the reaction such as temperature and concentration.

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Comment on the simulation.

- Although a relatively simple purely statistical simulation, this exercise reflects real chemical reactions in several ways. If we take all of the groups' results for a single turn, the fact that not every group produces a product molecule in a single turn reflects the fact that not every collision between molecules will be between reactants.
- We must break the Lego molecules and then reform them reflecting the breaking and making of bonds in the real reaction.
- We have to use our hands to break the bricks apart, reflecting the need for an energy of activation.
- We also need a correct orientation, otherwise the product molecules are not made properly.
- With the rule of the backward reaction when two product molecules encounter each other, we actually reach a dynamic equilibrium after about 25-30 turns.
- At a macro level, it seems that the concentration of product remains constant, on the micro-scale we can see that the reaction is going backwards and forwards at the same rate.

$$\text{rate} = \frac{\Delta[P]}{\Delta t} = \frac{1200-1000}{23.5-8} = 12.9 \text{ products/sec}$$

Tangent to curve at 20s

No. of product molecules

time (s)

The rate of reaction

In a reaction we have reactants which are converted to products. In some cases we have a single reactant which under the conditions of the reaction converts to products; in other cases we have mixtures of two reactants (for example the different coloured Lego bricks) which combine to form products. In rarer cases we may have three different components reacting together.

A very general way of describing the reaction between two reactants A and B is:

A	**+**	**B**	**→**	**C**
Reactant A		Reactant B		Product
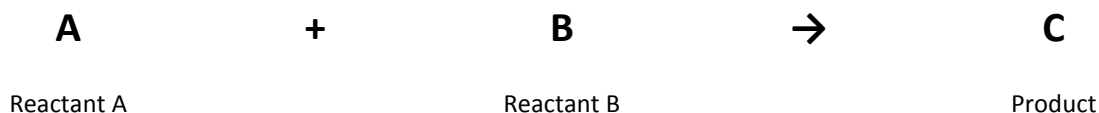
An example of this type of reaction is the reaction of hydrogen with iodine to form hydrogen iodide. When hydrogen gas and iodine vapour are heated together, they form hydrogen iodide (this is a classic experiment used in the study of equilibrium).

H_2	**+**	**I_2**	**→**	**2HI**
Reactant A		Reactant B		Product
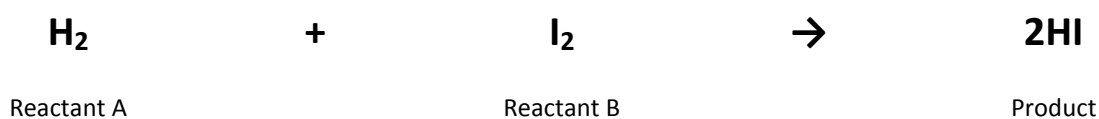
The rate of the reaction, that is how fast it goes, can be described mathematically with a rate equation.

The rate is related to the concentrations of [A] and [B] by the formula:

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

where k is called the rate constant and n and m are exponents which describe the relationship between the concentration and the observed rate. We can only find n and m by experiment – they do not correspond to the stoichiometric coefficients in the balanced equation.

The subject of kinetics deals with how to find out these values in order to understand the reaction in more detail.

Finding the order and the rate constant of a reaction: the thiosulphate reaction.

(Adapted from Flinn scientific: https://www.flinnsci.com/media/622120/91860.pdf *and* http://www.chemguide.co.uk/physical/basicrates/orders.html)

The purpose of this demonstration is to investigate the effect of sodium thiosulphate concentration on the rate of reaction of sodium thiosulphate with hydrochloric acid. The reaction, which produces solid sulphur, will be followed by measuring the time needed for the reaction mixture to become opaque. The results will be analyzed graphically to determine the order of reaction— the mathematical relationship between the reactant concentration and the rate.

Materials per group

• Hydrochloric acid solution, HCl, 2 M, 25 mL • Graduated cylinder, 10-mL, • Graduated cylinder 25ml or 50ml • Sodium thiosulphate solution, Na2S2O3, 0.15 M, 75 mL • Distilled or deionized water	• Permanent marker • Beaker, 100-mL, • Stirring rod • Stopwatch or timer/ smartphone • Sheet of white paper

Safety Precautions and disposal
> *Hydrochloric acid solution is corrosive to eyes and skin.*
> *Sodium thiosulphate solution is an irritant.*
> *The reaction generates sulphur dioxide gas, which is a skin and eye irritant. USE A WELL-VENTILATED LAB!*
> *Avoid contact of chemicals with eyes and skin.*
> *Wear chemical splash goggles, temperature-resistant gloves, and chemical-resistant lab coat.*

Procedure
1. Take a 100ml beaker and make sure that the bottom is clean.
2. Draw a large X on a piece of white paper, place the beaker on top of this.
3. Using separate graduated cylinders for the solution and water, measure and add the required amounts of 0.15M sodium thiosulphate and distilled water to the beaker according to the experiment number. Be as precise as possible.
4. Measure the 2M hydrochloric acid in a 10-mL graduated cylinder.
5. Carefully add the HCl to the sodium thiosulphate solution all in one go. Stir the solution once with a stirring rod and immediately start timing.
6. Stop timing when the black "X" is no longer visible. Record the reaction time in seconds in the data table.
7. Clean the beaker and carry out the next experiment with a different concentration of thiosulphate solution (see table).

Results
- Collect the data IN THE TABLE PROVIDED
- Calculate Ln(concentration) and 1/concentration
- Plot graphs of concentration, Ln(conc) and 1/conc against time
- Describe the shapes of the graphs obtained.

1) Effect of concentration of thiosulphate

Experiment	1	2	3	4	5
0.15 M Na2S2O3	25ml	20ml	15ml	10ml	5ml
Distilled Water	0ml	5ml	10ml	15ml	20ml
HCl (2M)	10ml	10ml	10ml	10ml	10ml
Total volume					
Concentration thiosulphate (M)					
Reaction time t (s)					

2) Effect of concentration of HCl

Experiment	1	2	3	4	5
0.15M Na2S2O3	15ml	15ml	15ml	15ml	15ml
2M HCl	10	5ml	2.5ml	1.0ml	0.5ml
Distilled Water	10ml	15ml	17.5ml	19ml	19.5ml
Total volume					
Concentration acid (M)					
Reaction time t (s)					

Order of reaction.

The reaction order can be a difficult concept but the key point is that *it is always found by experiment* – you cannot know it simply by writing the equation. Why? Because it is linked to how the reaction actually takes place and in particular which step of the reaction is the slowest.

Assume that from the reaction you have carried out you find that rate is directly proportional to the concentration of the reactant A. As you double the concentration of A, the rate doubles. We can write this as:

$$\text{rate} \quad \alpha \quad [A]$$

Where the square brackets is the usual way for writing concentration (moles per litre). If we add a constant of proportionality, k we remove the proportionality sign. "k" is the rate constant.
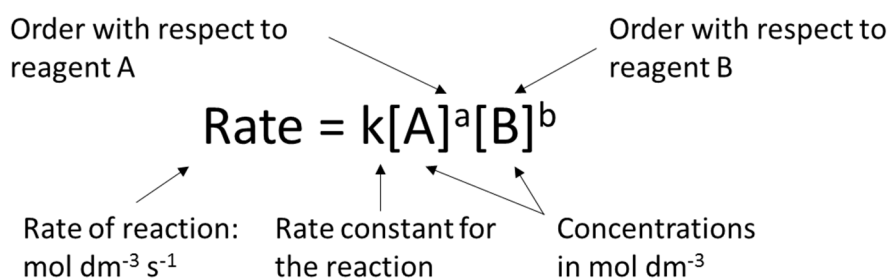
Maybe though our reaction wasn't so easy. Let's say we found out that the rate was proportional to the square of the concentration of A. In this case, as the concentration doubles, the rate increases by four times, ie the square. In this case we can write:

$$\text{rate} \quad \alpha \quad [A]^2$$

$$\text{rate} \quad = \quad k[A]^2$$

We can generalise this to a reaction between A and B in which we use the exponents (the powers) of the concentrations to show how they affect the rate of the reaction. A consequence of this is that the order is always expressed in terms of a particular reactant because the powers don't have to be the same for each reactant. These powers are called the *order of reaction.* For example, "the reaction is second order in A".

The rate of a reaction can be summarised with this relationship:

Order with respect to reagent A

Order with respect to reagent B

$$\text{Rate} = k[A]^a[B]^b$$

Rate of reaction: mol dm^{-3} s^{-1}

Rate constant for the reaction

Concentrations in mol dm^{-3}

It is possible to for a and b to be fractions and for a reaction to be zero order in a reactant. In this case, this term disappears from the rate equation. We can also describe the overall order of a reaction by summing the powers. For example: "the reaction is overall second order".

Experimentally, we find the order with respect to a particular reactant by drawing graphs. We are looking for straight lines which come from the mathematical integration of the rate equation.

If the relationship is:

| Zero order	Concentration vs time	Straight line, gradient = -k
First order	Log(concentration) vs time	Straight line, gradient = -k
Second order	1/concentration vs time	Straight line, gradient = k

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Some examples:

A + B = products

Rate law found by experiment	Comment
$\text{rate} = k[A][B]$	the order of reaction with respect to both A and B is 1. The overall order of reaction is 2 - found by adding up the individual orders.
$\text{rate} = k[B]^2$	This reaction is zero order with respect to A because the concentration of A doesn't affect the rate of the reaction. The order with respect to B is 2 - it's a second order reaction with respect to B. The reaction is also second order overall (because 0 + 2 = 2).
$\text{rate} = k[A]$	This reaction is first order with respect to A and zero order with respect to B, because the concentration of B doesn't affect the rate of the reaction (remember $[B]^0 = 1$). The reaction is first order overall (because 1 + 0 = 1)

Even if you have a different number of reactants, the concentration of each one will be represented in the rate equation raised to a power (remember, it could be 0, 1....). The overall order is found by summing them.
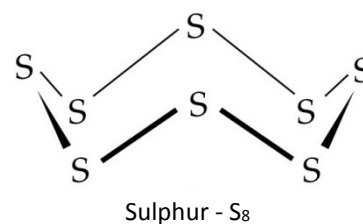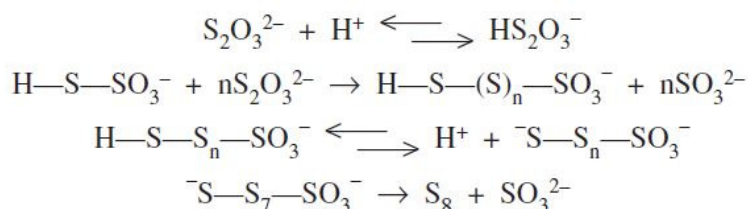
Discussion – finding the order of reaction.
Sodium thiosulfate reacts with hydrochloric acid to form sulfur and sulfur dioxide (Equation 1).

$$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow S(s) + SO_2(g) + 2NaCl(aq) \quad \textit{Equation 1}$$
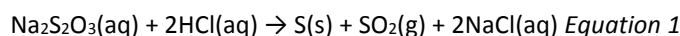
The kinetics of the reaction can be analyzed by the graphical method plotting the concentration of thiosulphate against time.
- A simple plot of concentration versus time gives a curved line, which levels off as it approaches the x-axis—the reaction slows down as the reactant concentration decreases. The rate of a reaction is inversely proportional to reaction time.
- A plot of the log of concentration against time also gives a curve.
- A plot of 1/concentration versus time gives a straight line. The reaction appears to be second order in thiosulfate, at least in this range of concentrations.
- When lower concentrations of sodium thiosulfate are used, the rate law does not appear to be as simple as predicted in this experiment. The order with respect to the acid is more complex and changes significantly with concentration.
- The reaction time is more difficult to measure at lower concentrations because the onset of turbidity is more gradual.
- Both the overall chemical equation and the mechanism for the decomposition of sodium thiosulfate are more complex than suggested by Equation 1. The reaction is acid-catalyzed, which means that the acid concentration must have some bearing on the rate in terms of producing an equilibrium concentration of $HS_2O_3^-$ ions, the $HS_2O_3^-$ ion is a reactive intermediate, reacting further with additional $S_2O_3^{2-}$ ions to produce polymeric ions containing multiple S atoms. When the chain of S atoms in a polymeric ion becomes long enough, it "closes" in on itself to form a ring of elemental sulfur (S_8).

$$S_2O_3^{2-} + H^+ \rightleftharpoons HS_2O_3^-$$
$$H-S-SO_3^- + nS_2O_3^{2-} \rightarrow H-S-(S)_n-SO_3^- + nSO_3^{2-}$$
$$H-S-S_n-SO_3^- \rightleftharpoons H^+ + {}^-S-S_n-SO_3^-$$
$$^-S-S_7-SO_3^- \rightarrow S_8 + SO_3^{2-}$$

Sulphur - S_8

What happens if we heat or cool down a reaction?

- Think about what happens to a cup of milk if you leave it out of the fridge in the summer or in the winter.
- Think about lizards and insects in the winter or in the summer.

We know from experience that if we heat a reaction, it will go faster. Why does this happen?
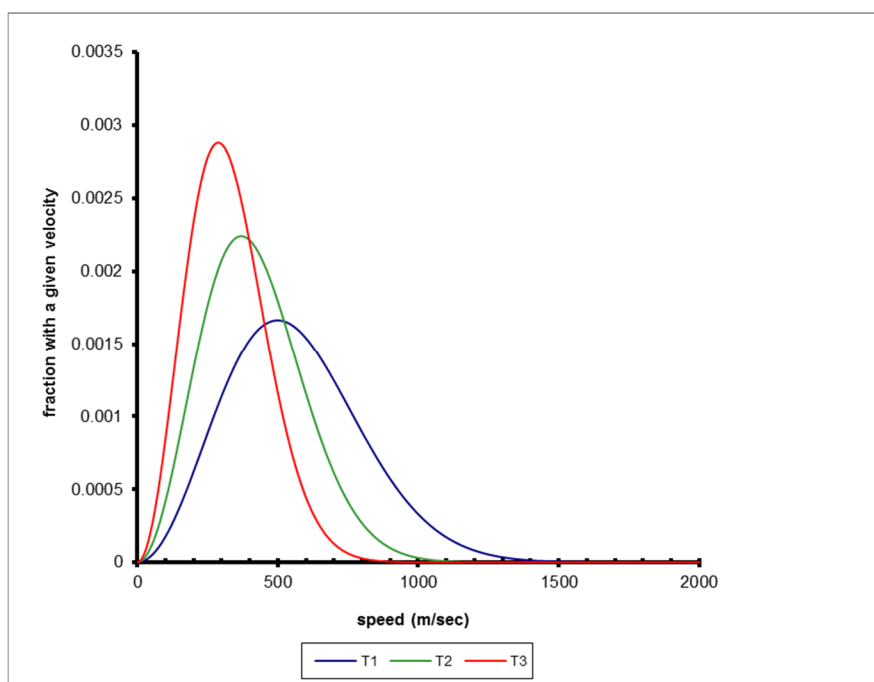
Kinetic energy of molecules and temperature.

- If we have a collection of gas molecules in a jar, do they all have the same energy? If they aren't the same, what do you think a distribution of their energies would look like?

Collisions

First of all, for the reaction to occur, we must have a collision. The two components must come together. However, a collision alone is not sufficient. A chemical reaction involves making and breaking bonds. We must break the bonds in the reactants and we must form new bonds in the products. This means that the reactants must have enough energy to break these bonds. This allows the reaction to take place.

In the late 1800s, Boltzmann showed that in a *population* of molecules, the molecules have different energies that are distributed over a range which is described by the Maxwell-Boltzmann law. This law relates the number of molecules with a given energy to the kinetic energy and is a probability distribution. This means that it shows the probability that a molecule possesses a particular energy. Since the area under the curve must always be 1, the curve must become broader and flatter as we increase the temperature and the average energy increases (moves to the right).


What does this mean?

The graphs show that as we increase the temperature, the average energy of the molecules increases which means that we have more molecules with higher energy. This point is very important when we are talking about the kinetics of a reaction because it means that at higher temperatures there is more energy available to break and make bonds.

Activation energy

We said that for a reaction to occur we need the molecules involved to come into contact with each other via a collision, but we also need them to have a certain amount of energy to break and make bonds. The energy required to do this is the so-called *activation energy.*
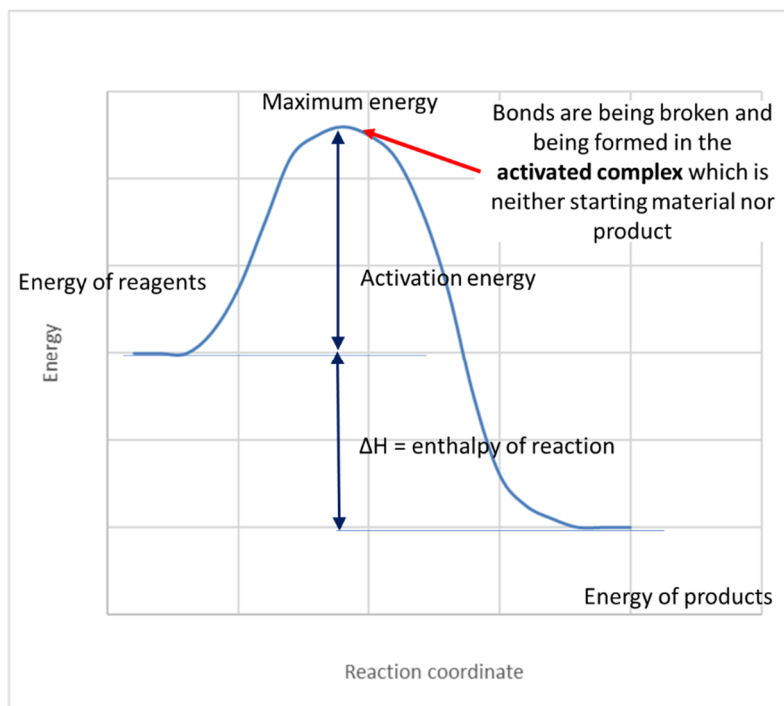
If we look at how the energies of the molecules change as the reaction proceeds and plot a graph, we find that there is a hump. This is the activation energy which is characteristic of a particular reaction. This is the energy needed to get the reaction to go.

Looking closely at the graph we can see that there is an activation energy for the forward reaction as well as one for the reverse reaction and the peak of the energy curve corresponds to the so-called activated complex. This is the point of the reaction in which the arrangement of molecules is neither reactants nor products: they form a complex in which the formation and breaking of bonds is at a maximum.

Note how the thermodynamic value of enthalpy change ΔH does not take into account the activation energy, it refers to the relative energy values of the products and reactants. This is why thermodynamics and kinetics tell us different things about the reaction and why thermodynamics cannot tell us whether a reaction will actually occur.

Where does the energy for a reaction come from?

The energy needed for a reaction is usually supplied as *thermal energy*, as we heat the reaction up. This increases the average kinetic energy of the molecules and the number of collisions and hence the reaction goes faster. Remember, we need to break bonds and make bonds. This energy will also break the weaker bonds within one or more reagents to give reactive species such as radicals. We can also supply energy in the form of light, typically short wave high energy light such as UV (has anyone had any white fillings at the dentist recently?) or electrochemical energy where we supply a voltage difference to promote the reaction by breaking weak bonds in reagents. Many chemical reactions start with the formation of radicals:

$$H_2 \quad + \quad Cl_2 \quad \rightarrow \quad 2HCl$$

The Arrhenius law

By 1890 it was well known that the reaction rate consistently increased with temperature. Experimentally it had been shown that for many reactions, an increase of 10°C caused a doubling of reaction rate. This relationship is so well-known that it has been used to estimate temperature based on stridulating crickets in California (see *http://www.dartmouth.edu/~genchem/0102/spring/6winn/cricket.html*). Since the noise that crickets make is dependent on chemical reactions which depend on temperature, the frequency of the stridulation can be used to estimate the ambient temperature!

Arrhenius described the connection between temperature and rate in mathematical terms by bringing together the work of Maxwell and Boltzmann on the kinetic energies of gas molecules and the concept of activation energy. His theory is summarised in the following equation which looks fearsome but which is actually not so hard to understand:

$$k = A\exp^{-(Ea/RT)}$$

- where k is the rate constant
- A is a constant factor for the reaction being studied
- T is the temperature
- R is the gas constant (8.314×10^{-3} kJmol^{-1})
- and RT is the *average kinetic energy*
- E_a is the activation energy

If we ignore A for the moment, we just need to know that it is a constant, the first thing that we notice is that the relationship is a negative exponential. That means that something is decaying rapidly as the value of Ea/RT gets bigger.

From the equation we can see that the size of the rate constant k is related to the ratio Ea/RT. As this ratio gets bigger, the rate gets smaller.

Let's think about this a minute. At a given temperature, as the activation energy increases (that hump on the graph gets bigger!) the rate constant must decrease. This makes sense because as the barrier to reaction increases, fewer molecules have the energy required to complete the reaction.

If we plot a graph of the rate constant against the activation energy, we can see that it decreases rapidly as the activation energy increases. Since this is an exponential relationship, even small changes have a very large effect. If we do this at two different temperatures, we can see that the rate constant is higher at higher temperatures for a given activation energy.
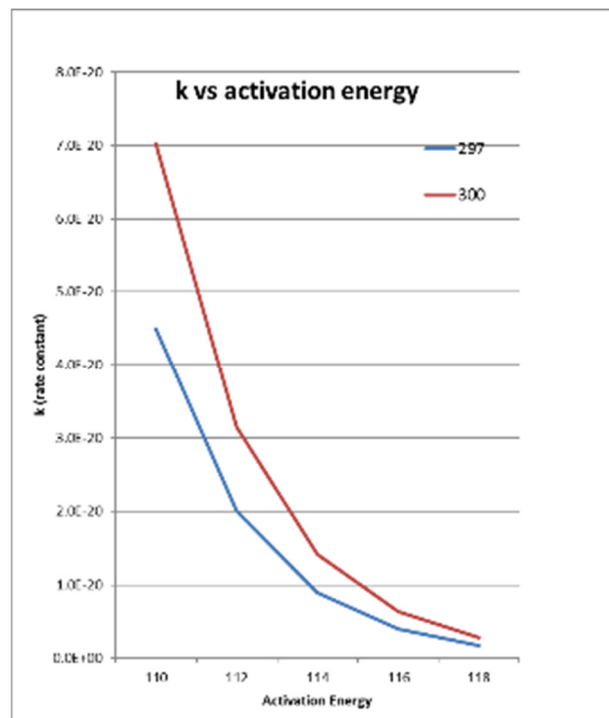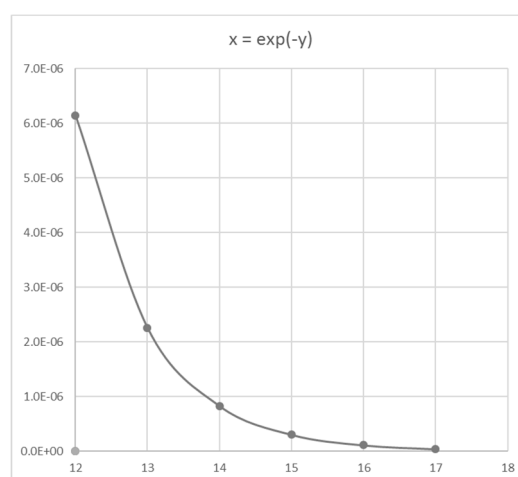
Estimating the activation energy for the thiosulphate reaction.

We can use a variation of the thiosulphate reaction to estimate the activation energy using the Arrhenius equation (for a very clear explanation see: http://www.chem1.com/acad/webtext/dynamics/dynamics-3.html). The brilliance of the Arrhenius equation is that it connects the rate constant to the activation energy meaning that if we find the rate as a function of temperature, we can measure the activation energy.

$$k = A \exp^{(-E_a/RT)}$$

Where k = rate constant; Ea = activation energy; T = temperature; R = gas constant

Quite simply we will measure the rate constant at two different temperatures and together with our data from the room temperature laboratory, we will use Arrhenius's law to estimate the activation energy.

Experiment

Carry out the thiosulphate experiment exactly as you did in the previous laboratory except *you must change the temperature.*

Cold bath (e.g. T= 10oC)

Use a container such as a larger beaker or ice-cream tub and add some cold water and then ice to cool the water to 10oC. Keep at this temperature by adding ice occasionally and stirring.

DO NOT MIX THE REAGENTS YET!!! Place the flask with the thiosulphate solution and the test-tube containing the HCl in the cold bath and let them sit for several minutes BEFORE STARTING THE REACTION (this is to make sure that they are at the right temperature).

After several minutes add the acid to the thiosulphate and start the timer.

Warm bath (e.g. 30oC)

DO NOT MIX THE REAGENTS YET!!! Using a thermostatic bath, warm the thiosulphate and HCl to the required temperature for a few minutes. After several minutes, add the HCl to the thiosulphate and start the timer as normal.
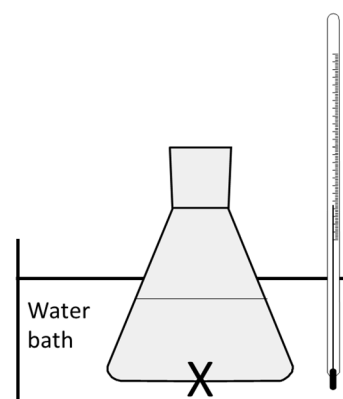
X on the bottom of the flask or under the beaker containing the water bath.

Experiment	1	2	3	4	5
0.15 M Na2S2O3	25ml	20ml	15ml	10ml	5ml
Distilled Water	0ml	5ml	10ml	15ml	20ml
Total volume					
Concentration thiosulphate (M)					
Reaction time t (s)					

NOTE:

- Low temperature reaction: the reaction may be too slow at the lowest concentration. Start at the highest concentration.
- High temperature reaction: the reaction may be too fast at the highest concentration. Start at the lowest concentration.

Data.

1) Use the data for your temperature to calculate the rate constant of each reaction at each temperature from the graph of 1/concentration vs time. Remember the units! (we are assuming 2nd order).
2) Then plot a graph of the log of the rate constants against 1/T. T should be measured in Kelvin and pay attention to the units of R, the gas constant. The gradient of the line gives a measure of Ea/R.

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Measuring the activation energy

With data from experiments in which we measure the reaction rate at different temperatures we can estimate the value of the activation energy graphically. The first step is to make the relationship linear. We do this by taking the natural log of each side of the equation:

$$k = A\exp^{-(Ea/RT)}$$

$$\ln k = \ln A - (Ea/RT)$$

$$\ln k = \ln A - (Ea/R). \, 1/T$$

$$y = c + mx$$

Therefore, if we draw a graph of lnk against 1/T, we will find that the line we obtain has a gradient which is the activation energy divided by the gas constant R (i.e. Ea / 8.314).

Data from the thiosulphate reaction.

T(°C)	T(K)	1/T	k*	Lnk
10	283	0.003534	0.17	-1.78748
20	293	0.003413	0.25	-1.37498
31.3	304.3	0.003286	0.41	-0.88971

measured in separate experiments (see graph below) and adjusted for concentration mol^{-1}
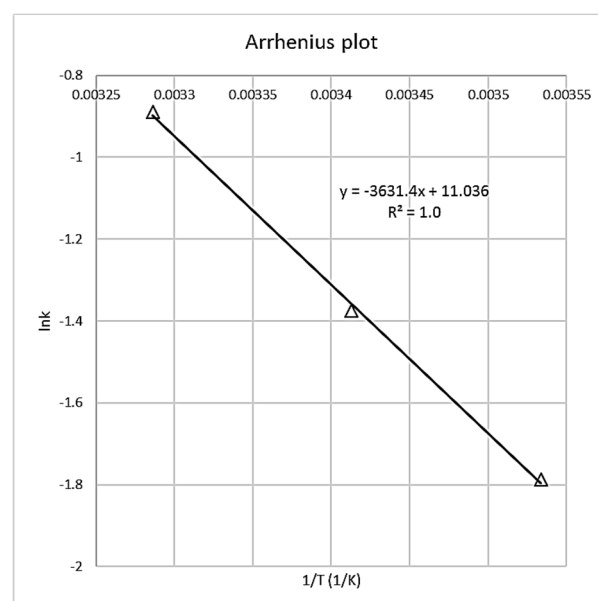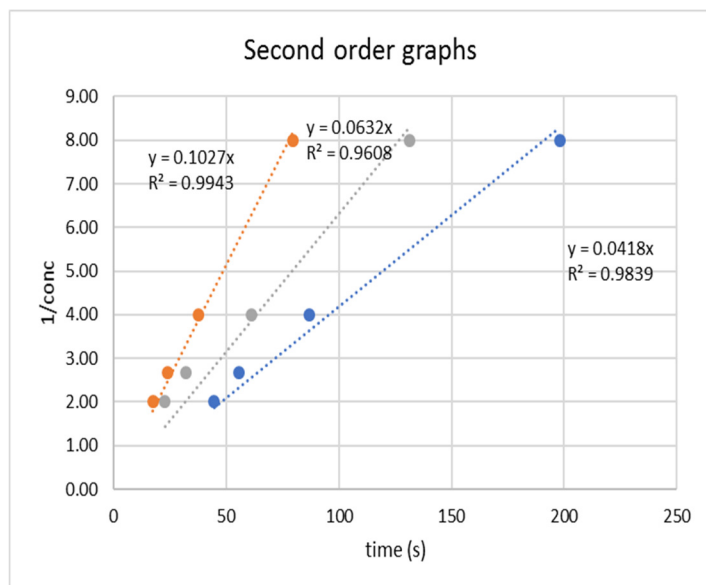
Using estimates of the rate constant measured at three different temperatures (derived from graphs of 1/[thiosulphate] vs time), the graph of lnk against 1/T shows a straight line. Working this in excel and applying linear regression, we find the following values:

Gradient: -3631.4 = -Ea/8.314 (Jmol^{-1}); therefore: activation energy Ea = 31 kJmol^{-1}

An estimate of the pre-exponential factor A: LnA = 11.036, therefore A = 62 x 10^3 (which is quite small)

What about A?

A is the part of the equation that takes account of the collisions. Remember the Lego game? To make the product, the pieces must be brought together: we must break the old bonds and create the new ones. The pieces in effect must

collide but if you look at the bricks it is quite clear that not all of the collisions can actually give rise to the product because the collision *must be in the correct orientation*.

*(http://www.thinctanc.co.uk/photography/miscellaneous.html)*

A is the factor which takes account of two aspects of the collision: how often molecules collide (frequency of collision) and the orientation. A is usually worked out from experiments and is considered to be constant for a particular reaction. The factors involved in orientation of the collision include the shape of the molecule and are called *steric* factors. The more complicated the structure of the molecule, the more important the steric factors are in determining whether the reaction can proceed.

A can be considered as the maximum rate of reaction if the activation energy is zero or if all of the molecules have energies that are higher than the activation energy (this doesn't happen often!).

Multistep reactions: the catalytic decomposition of hydrogen peroxide with potassium iodide.

We mentioned that the thiosulphate reaction is rather complicated and since the product is formed of molecules of S_8, it clearly can't occur in a single step. Another example of a multistep reaction is the catalytic decomposition of hydrogen peroxide with potassium iodide. In this case, it is possible to observe an initiation period in which the first reaction (which is quite slow) leads to the formation of an equilibrium concentration of a reactive intermediate that then reacts rapidly to form the product.

Hydrogen peroxide decomposes in the presence of potassium iodide (KI) producing gas and releasing heat. Overall the reaction follows this equation:

$$2H_2O_2 \rightarrow 2H_2O + O_2$$
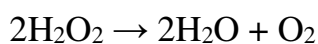
Qn: Why doesn't the KI appear in the equation?

Qn: How would you show that the gas produced is oxygen?

Measuring the speed of *foam*

Materials and equipment

|	
H_2O_2 in different concentrations (see instructor) Solution of washing up liquid and water 1:1 Potassium iodide (KI) solution (0.3g KI / 1ml water) Measuring cylinder	Plastic basin Camera or smartphone

Procedure
- Put the cylinder in the basin (to collect any solution which escapes)
- Put 10ml of H2O2 solution in the cylinder and make a note of the concentration.
- Add 1ml of washing up liquid solution to the cylinder
- Add 1ml of KI solution to the cylinder and mix it rapidly.
- Using the chronometer function of a smartphone or a watch, measure the volume of the foam with time recording the data in the table. (It is best to take a video!)
- Note any colour changes (photograph).

Repeat the experiment using different concentrations of hydrogen peroxide (see the table, ask the teacher) You could also study the effect of temperature and amount of KI. How would you do this?

Working with the data

- Draw graphs of the amount of oxygen produced vs time for the different concentrations of H_2O_2
- Describe the graphs.
- What can you say about the rate of the reaction?
- Can you measure the rate at the start of the reaction from the graphs?

A more complex story.

$$H_2O_2 + 2I^- \rightarrow 2OH^- + I_2$$
$$2OH^- + I_2 \rightarrow H_2O + I^- + IO^-$$
$$H_2O_2 + IO^- \rightarrow H_2O + O_2 + I^-$$
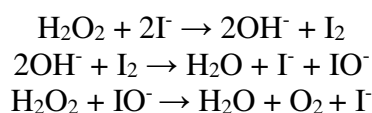
The reaction is more complex than the stoichiometric equation would have us believe: the iodide is clearly doing something (otherwise why would we add it?). The above series of reactions are thought to be responsible for the overall result.

- ❖ Are there any additional clues from the experiment that you did that could support this hypothesis?
- ❖ Are there other experiments that you could do to investigate this further?

RESULTS

Conditions: 10ml H_2O_2 solution (NOTE THE CONCENTRATION); 1ml detergent; 1ml of 0.3g/ml KI

Time	H2O2 concentration		
seconds	Volume of foam	Volume of foam	Volume of foam
0			
10			
20			
30			
40			
50			
60			
70			
80			
90			
100			
110			
120			
150			
180			
210			
240			
270			
300			
330			
360			
420			
480			
540			
600			