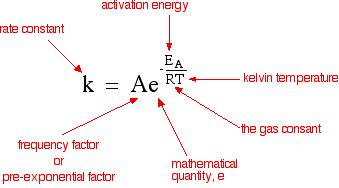
**NAME** ............................................ Chemistry Class ....................

Student Number ……….

Chemical Kinetics II

**Topic 9**

5. understand the role of catalysts in providing alternative reaction routes of lower activation energy

6. be able to draw the reaction profiles for uncatalysed and catalysed reactions

7. be able to interpret the action of a catalyst in terms of a qualitative understanding of the Maxwell-Boltzmann distribution of molecular energies

8. understand the use of a solid (heterogeneous) catalyst for industrial reactions, in the gas phase, in terms of providing a surface for the reaction

9. understand the economic benefits of the use of catalysts in industrial reactions

**Topic 16**

8. understand how to:

i obtain data to calculate the order with respect to the reactants (and the hydrogen ion) in the acid-catalysed iodination of propanone

ii use these data to make predictions about species involved in the rate-determining step

iii deduce a possible mechanism for the reaction

9. be able to deduce a rate-determining step from a rate equation and vice versa

10. be able to deduce a reaction mechanism, using knowledge from a rate equation and the stoichiometric equation for a reaction

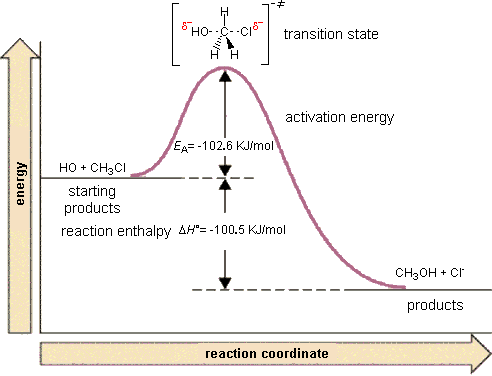
11. understand that knowledge of the rate equations for the hydrolysis of halogenoalkanes can be used to provide evidence for SN1 or SN2 mechanisms for tertiary and primary halogenoalkane hydrolysis

12. be able to use graphical methods to find the activation energy for a reaction from experimental data

*The Arrhenius equation will be given if needed.*

***CORE PRACTICAL 13a and 13b: Rates of reaction: Following the rate of the iodine-propanone reaction by a titrimetric method and investigating a ‘clock reaction’ (Harcourt-Esson, iodine clock)***

***CORE PRACTICAL 14: Finding the activation energy of a reaction***



**Chemical Kinetics and Rates of reactions**

****The first of the Kinetics packs you learnt about following the rate of a reaction

**What to do if you get stuck**

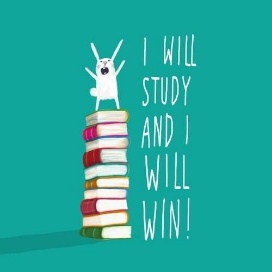
If you are struggling with this topic here are a few ideas you can try

1. Read Chapter 16 in the 2nd Chemistry textbook

2) Try some of the worksheets in the folder on GOL

3) Come to clinic on Friday 1-2pm and/or Tuesday 4.15-5pm for some 1-2-1 support from your teachers.

**What to do if you have time left over**

[](https://twitter.com/WeTeachRevision/status/438710135951593472/photo/1/large)

If you have finished all the set work from your teacher but you still have some of your 5 hours left over here are some things to try

3) Read and make notes on Chapter 9 in ‘Why Chemical reactions Happen’ copies are in the library

**Core Practical 14 Finding the activation energy of a reaction CPAC 2c, 4a, 5a**

|  |  |
| --- | --- |
| **CPAC 2:**  **Applies investigative approaches and methods when using instruments and equipment** | 1. Identifies and controls significant quantitative variables where applicable, and plans approaches to take account of variables that cannot readily be controlled. |
| **CPAC 4:**  **Makes and records observations** | 1. Makes accurate observations relevant to the experimental or investigative procedure. |
| **CPAC 5:**  **Researches, references and reports** | 1. Uses appropriate software and/or tools to process data, carry out research and report findings. |

MCj04247820000[1]References:- Edexcel A-level Chemistry Book 1 page 262 onwards

Edexcel A-level Chemistry Book 2 page 147 onwards

Factsheets

|  |  |
| --- | --- |
| 43 | Rate Equations, Orders and Constants |
| 44 | Rate Expression Orders and Experimental procedures |

**The rate determining step**

Ref Edexcel A-level Chemistry Book 2 pages 160-161

Cann and Hughes p217 - 218

There is plenty of evidence that the gas phase reaction between a hydrogen atom and bromine

H + Br2 → HBr + Br

Takes place exactly as written i.e. the products are formed from a single encounter between a hydrogen atom and Br2. This is an example of an **elementary** reaction, i.e. the reaction takes place in a single encounter, as the equation is written.

However most of the reactions we study are not so straight forward. They do not proceed by one simple step, rather they are a series of reactions. For example the oxidation of glucose has the overall equation:-

C6H12O6 + 6O2 → 6CO2+ 6H2O

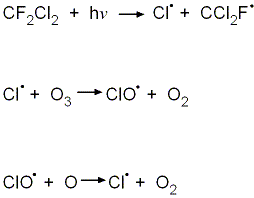
However thinking back to collision theory it should be self evident that this reaction where, if it were to work as written, would involve the simultaneous collision of 7 molecules all with the correct energy and in the correct orientation is not possible.

This is an example of a **complex** reaction. Complex reactions take place by a set of elementary reactions (single encounter), called a **mechanism;** typically such a mechanism will involve the generation of molecules called intermediates, which do not appear in the stoichiometric equation.

For example the catalytic decomposition of ozone into oxygen has the stoichiometry:-

2O3 → 3O2

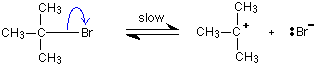
This overall reaction has a mechanism involving 3 elementary steps

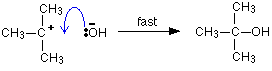


Determining this mechanism is not simple task. It is not possible to deduce from looking at the equation what the mechanism will look like. It must be determined by experiment. By studying the rate of reaction we can begin to understand the mechanism.

The overall rate of a reaction (the one which you would measure if you did some experiments) is controlled by the rate of the slowest step in the mechanism.

Here is a mechanism we have already looked at. How do we know that it works like this?





By doing rate of reaction experiments, you find this rate equation:

rate = k[(CH3)3C-Br]

The reaction is **first order** with respect to the organic compound, and **zero order** with respect to the hydroxide ions. The concentration of the hydroxide ions is not affecting the overall rate of the reaction. This is consistent with an SN1 mechanism, i.e. only 1 species in the slow step.

If the hydroxide ions were taking part in the slow step of the reaction, increasing their concentration would speed the reaction up. Since their concentration does not seem to matter, they must be **taking part in a later fast step.**

Increasing the concentration of the hydroxide ions will speed up the fast step, but that will not have a noticeable effect on the overall rate of the reaction. That is governed by the speed of the slow step **the rate determining step**.

Rate determining step

The slowest stage in a reaction and is specified by the overall rate equation for the reaction.

**1.** (b) Two gases, A and B, react according to the equation

A + 3B AB3

A series of experiments carried out at 298 K gave the following results.

|  |  |  |  |
| --- | --- | --- | --- |
| Experiment | Initial concentration of A  / mol dm-3 | Initial concentration of B  / mol dm-3 | Initial rate of  reaction  /mol dm-3 min-1 |
| 1 | 0.100 | 0.100 | 0.00200 |
| 2 | 0.100 | 0.200 | 0.00798 |
| 3 | 0.200 | 0.100 | 0.00399 |

(i) State the order of reaction with respect to each of the reactants. Justify your answer.

Reactant A ……………………………………………………………………

................................................................................................................................

................................................................................................................................

Reactant B ........................................................................................................

................................................................................................................................

................................................................................................................................

**(3)**

(ii) Write the rate equation for the reaction between A and B.

................................................................................................................................

**(1)**

(iii) Use the experimental data from **Experiment 1** to calculate the rate constant, including units.

**(2)**

(iv) Suggest a possible mechanism for the reaction between A and B, leading to the formation of AB3. Identify the rate-determining step.

................................................................................................................................

................................................................................................................................

................................................................................................................................

................................................................................................................................

**(3)**

Leave blank

June 2006 CH5 Q3

Look back at the reaction between iodine and propanone in the previous pack. This is a favourite of examiners!

I2 (aq) + CH3COCH3(aq) CH2ICOCH3(aq) + HI (aq)

The reaction was found to be 1st order with respect to propanone, 1st order with respect to H+ (aq) and 0 order with respect to iodine.

Does iodine take part in the rate determining step? ………………

The **rate determining step** therefore involves a reaction between …….............and ……………

New exercise

Give the equation or mechanism to show this slow step

So, from the orders of reaction we can gain a better understanding of the mechanism of the reaction.

**Non-integral orders**, i.e. values between 0, and 1 or between 1 and 2 etc; such reactions are either chain reactions or have more than one rate determining step.

From your mechanism pack you will be aware that there are 2 types of nucleophilic substitution. SN1 and SN2 reactions

**SN1 reactions**

These involve only 1 species in the slow rate determining step.

|  |
| --- |
| mesngif**SN1 reaction.** |
|  |

So this reaction is 1st order with respect to the organic molecule and 0 order with respect to the nucleophile.

The hydrolysis of the tertiary halogenoalkane 2-bromo-2-methylpropane is an SN1 reaction.

**Curly arrows**

They represent the movement of a pair of electrons, either a bond pair or a lone pair.

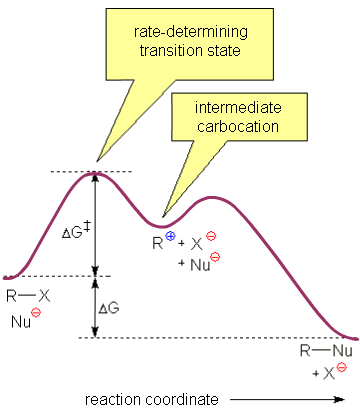
A (double headed) curly arrow must:-

Start from a bond (σ or π) or from a lone pair of electrons on an atom or ion

* Go to an atom to form an anion **or**
* Go towards an atom to form a bond

New exerciseDraw out the mechanism for this reaction indicating the rate determining step

Note that this reaction involves **an intermediate carbocation** being formed **Enthalpy level diagram of an SN1 reaction**



H

H

Enthalpy level kJ mol-1

New exerciseNotice the two energy peaks, the reaction has an intermediate stage.

Draw the energy level diagram for the previous reaction and **label each stage**

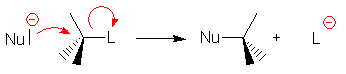
It is important not to confuse intermediates with transition state. **A transition state** is simply an arrangement of atoms that the reaction must pass through and is located at the maximum in the energy profile (see above), note that transition states are not real molecules, and hence cannot be isolated.

**Intermediates** are, like products and reactants, found at the minima in the energy profile. They are ‘real’ molecules, usually short lived and highly reactive, however they do not have the fleeting existence of transition states and in some cases can be isolated.

Write the rate equation for this SN1 reaction

**The SN2 reaction** by contrast involves 2 species in the rate determining step

In this reaction, the nucleophile (Nu) attacks the substrate from the side opposite to the leaving group (L).

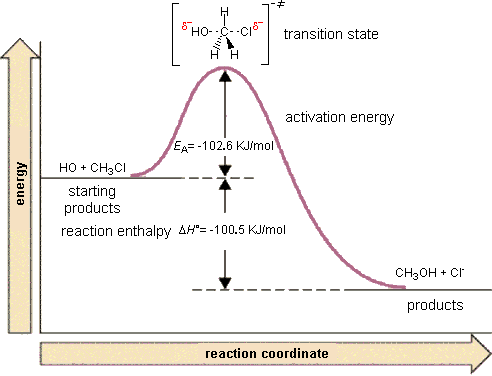


New exerciseSo a SN2 reaction is first order with respect of both the organic molecule and the nucleophile, as they are both involved in the rate determining step.

Write out the **mechanism** for this and **label each step.**

Write the rate equation for this SN2 reaction

Note that this reaction involves **a transition state** denoted by [ ], it cannot be isolated or observed, its existence is postulated from thermodynamic evidence only.



**Energy level diagrams**

**New exercise**Draw and label an energy level diagram for a reaction which is exothermic in the forward direction and proceeds through a transition state, (i.e. no intermediates are formed).

Label your diagram

If the reaction is exothermic in the forward direction, it is endothermic in the reverse direction.

Label the Ea for this reaction.

Using the diagram, explain what is happening at:-

1. The upward part of the curve……………………………………………

…………………………………………………………………………………..

……………………………………………………………………………………

1. At the peak……………………………………………………………………

……………………………………………………………………………………..

1. Downward part of the curve………………………………………………….

………………………………………………………………………………………

**Kinetic stability**

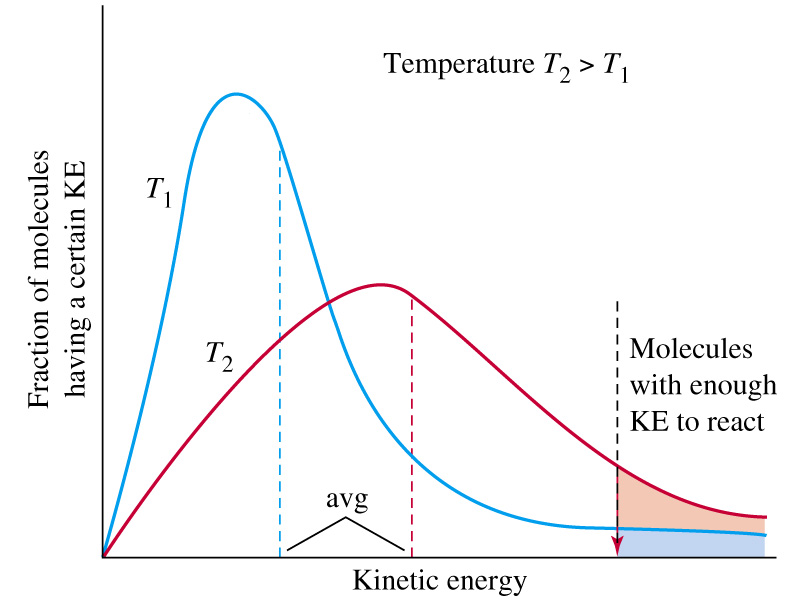
Reactions are **kinetically stable** when the activation energy is so large that there is almost no molecules with sufficient kinetic energy > Ea and so the reaction does not happen to an appreciable extent or happens too slowly to measure.

**Thermodynamic stability**

If the enthalpy change for a reaction is positive (H = +) the reactants are lower in energy than the products. The reactants are said to be **thermodynamically stable** with respect to the products; such reactions are unlikely to take place.

The rate of a reaction therefore depends on both the entropy change **and** the activation energy

**Maxwell –Boltzmann distribution of molecular energies**



Area under the curve represents number of molecules with enough KE to react

average KE

at T1 and T2

Why does the curve start at the origin? ............................................................................................

……………………………………………………………………………………………………...

Why does the curve not intersect the x axis at higher energies? ......................................................

……………………………………………………………………………………………………...

…………………………………………………………………………………………...................

What has happened to the number of molecule with energy equal to Ea as the temperature increases from T1 to T2? ..................................................................................................................

……………………………………………………………………………………………………..

As the temperature increases why does the rate increase…………………………………………..

…………………………………………………………………………………………...…………

………………………………………………………………………………...................................

…………………………………………………………………………………………...................

…………………………………………………………………………………………...................

On the graph draw a distribution curve for T3 where T3 is greater than T2

**Catalysts**

New exerciseOne factor we have not looked at yet is the effect of catalysts on the rate of reaction.

A Catalyst is defined as

**Types of catalyst**

Catalysts can be divided into three types:

1. **Homogeneous catalysts** – Here the catalyst is in the **same physical state as the reactants.** (usually liquids or in solution). The catalyst works by reacting with one of the reactants to form an **intermediate** or **activated complex** which then reacts with the other reagent to reform the catalyst. The catalyst often changes oxidation state in the process.

e.g Fe2+(aq) ions act as a catalyst for the oxidation of iodide ions by peroxidodisulfate ions:

S2O82-(aq) + 2I-(aq) 🡪 2SO42-(aq) + I2(s)

When the Fe2+ catalyst is present the following 2 **fast** steps are thought to take place:

S2O82-(aq) + 2Fe2+ (aq) 🡪 2SO42-(aq) + 2Fe3+(aq)

2Fe3+(aq) + 2I-(aq) 🡪 2Fe2+ (aq) + I2(s)

1. **Heterogeneous catalysts** – Here the catalyst is in a **different physical state to the reagents.** Normally the catalyst is a solid and the reactants are gases.

e.g. The Haber process for the manufacture of ammonia is catalysed by solid iron.

N2(g) + 3H2(g) 🡪 2NH3(g)

The solid catalyst has **active sites** which **adsorb** the gas molecules onto the **surface**. This **weakens the bonds** in the reacting molecules and **brings them** **together** so they are more likely to collide and react. The product molecules are then **released from the surface** of the catalyst so that it can be reused. Draw a diagram to show this process in the box below:

1. **Enzymes** – these are **biological catalysts**

e.g. catalase in the liver speeds up the decomposition of hydrogen peroxide in the blood

2H2O2(aq) 🡪 2H2O(l) + O2(g)

The enzyme has an **active site** of **specific shape** to bind the reactant (substrate). The reaction takes place on the active site where the **original bonds are weakened.** When the reaction has taken place the **product molecules leave the active site** so the enzyme can be re-used.

**Experimental study of catalysts**

Demonstration**Homogeneous catalyst with a visible intermediate** Classic Chemistry 1 demonstration

New exerciseA solution of potassium sodium tartrate (KNaC4H4O6.4H2O) (5g in 60 cm3) reacts very slowly with 20 cm3 of ’20 volume’ hydrogen peroxide solution (even when heated to 75oC) according to the ionic equation:

5H2O2(aq) + C4H4O62-(aq) 🡪 4CO2(g) + 6H2O(l) + 2OH- (aq)

State the formula of the catalyst added: …………………… Colour ………………………………

Observations when catalyst is added …………………………………………………………………

…………………………………………………………………………………………………………

…………………………………………………………………………………………………………

In what ways does this reaction show the typical behaviour of a homogeneous catalyst?

…………………………………………………………………………………………………………

…………………………………………………………………………………………………………

**2. Catalysts for the decomposition of H2O2** Classic Chemistry 58 demonstration

Equation…….…………………………………………………………………………………………

The following catalysts are added to 150cm3 of ’50 volume’ hydrogen peroxide in 5 large measuring cylinders to which a little washing up liquid has been added.

|  |  |
| --- | --- |
| **Catalyst** | **Observations** |
| 0.5g MnO2 |  |
| 0.5g PbO2 |  |
| 0.5g Fe2O3 |  |
| 1 cm3 liver (catalase) |  |
| 1 cm3 potato (catalase) |  |

Which catalyst increased the rate of reaction?

• the most? ......................................... the least? ...............................................

What three things do **all** catalysts have in common when considering their effect on a reaction?

…………………………………………………………………………………………………………

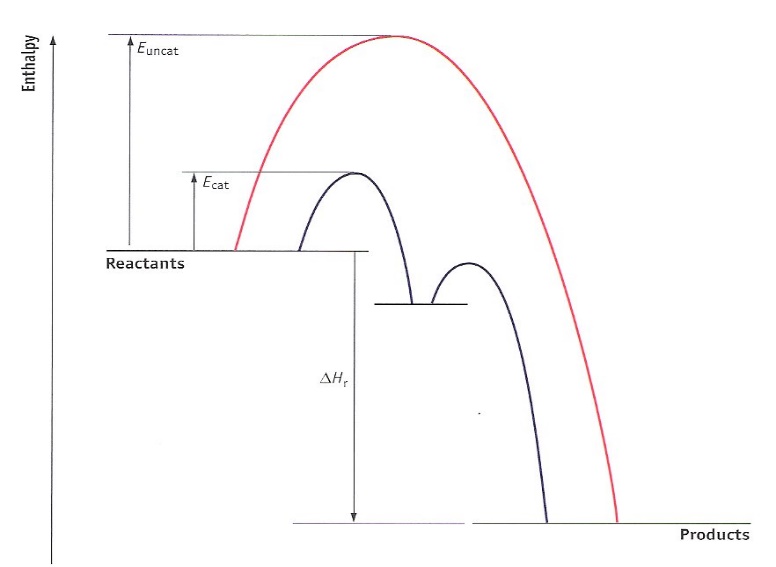
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**Reaction profiles to show the effect of catalysts**

Reaction profile for a catalysed and uncatalysed exothermic reaction.



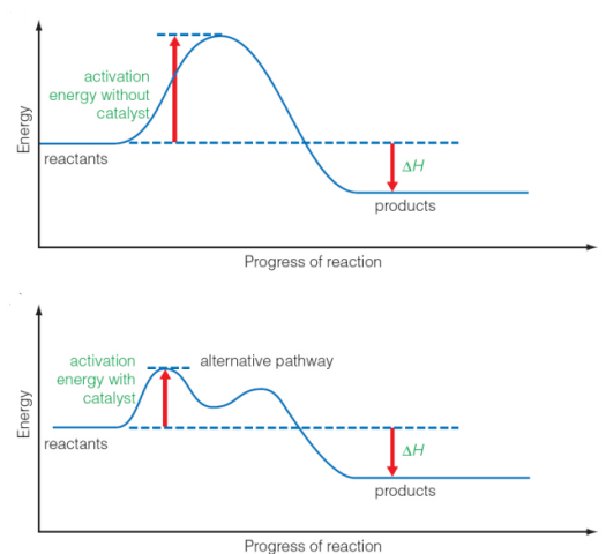
What do you notice about the activation energies of the uncatalysed and catalysed reactions?

………………………………………………………………………………………………………

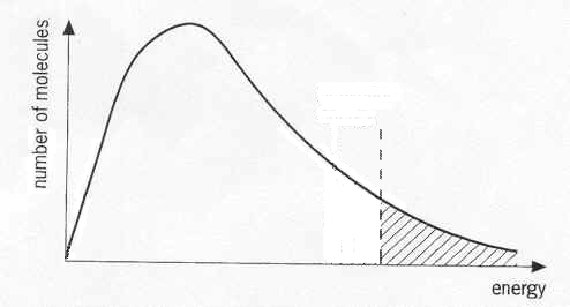
Label the diagram to show **the intermediate formed** between the catalyst and the reactants. It has some stability (it exists for long enough for you to see it in the previous demonstration). It will react again to reform the catalyst and form the products.

Draw a fully labelled reaction profile of an endothermic catalysed reaction

Another exothermic catalysed reaction:-

Label **the intermediate**.

**Maxwell Boltzmann diagrams and the effect of a catalyst**



On the diagram above:-

Label the EAuncat. Show the effect of adding a catalyst. If you are unsure about the effect of a catalyst, look back at the definition of a catalyst.

The lower value for the catalysed activation energy, Eacat., results in a greater proportion of molecules having more than the minimum energy required for a successful collisions.

Shade these in on your diagram.

Explain why adding a catalyst increase the rate of a reaction using the collision theory model

…………………………………………………………………………………………………………

…………………………………………………………………………………………………………

…………………………………………………………………………………………………………

…………………………………………………………………………………………………………

**The Arrhenius Equation**

Referring back to your work on kinetics you will remember that a temperature increase caused an increase in the rate of reaction. We are now in a position to investigate this further.

The Swedish chemist [Svante Arrhenius](http://en.wikipedia.org/wiki/Svante_Arrhenius) found that the rate constant depends on 3 factors

1. The orientation factor, particles must collide in the correct orientation for the reaction to occur. This is constant for a particular reaction
2. The activation energy. A large activation energy results in a small rate constant and hence a small rate of reaction.
3. The temperature

Thus we can derive the Arrhenius equation which relates rate constant, activation energy and temperature.

activation energy

rate constant

**-**EA

RT

temperature in Kelvin

k = Ae

the gas constant

Arrhenius constant or pre-exponential factor

mathematical quantity, e

You do not need to learn the equation only manipulate it. It would always be given to you in an exam

Ref: Cann and Hughes p221

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

It is useful as it explains why a **small increase in temperature has a large effect on the rate**

Another way of expressing the equation is by taking natural logs:-

ab = c

so loga c = b

see GOL

EA

RT

ln k = ln A -

= ln A -

Using this equation what is the relationship between the rate constant k, and the temperature T?

EA **.**  1

R T

……………………………………………………………………………………………………

……………………………………………………………………………………………………

Using this equation, what is the relationship between the rate constant k, and the activation energy?

……………………………………………………………………………………………………

……………………………………………………………………………………………………

By rearranging the equation ln k = -Ea **.** 1 + lnA

R T

it becomes the ? familiar y = m  **.** x + c

so we can use it to plot a graph:-

Plot of natural log of k against the reciprocal of temperature in K

Do **NOT start axes at zero**

Waste of space!

Use a **scale** that uses

**> ½ graph paper**

1/Temp in Kelvin

intercept = ln A

ln k

1/T K-1

intercept = ln A

dy

gradient = dy

dx

Natural log of

1 .

time for soln.

to turn blue

= -ve gradient

Slope = - Ea

R

dx

We can use the Arrhenius equation to calculate the activation energy (see p 13, 14 & 15) by following a reaction at different temperatures, calculating the rate constant for each experiment. As our change in concentration is constant for all the experiments, we can say that k = 1/time. Plotting ln k vs. 1/temp (in Kelvin) should give us a straight line. The gradient of the line = -Ea/R where R is a constant.

**Experiment - determining the activation energy for a reaction CPAC 2c, 4a, 5a**

The Arrhenius equation explains why a small increase in temperature causes a large increase in reaction rate

*Please note that you are not expected to remember the Arrhenius equation but you are required to be able to use it:-.*

Where

k=rate constant

A= Arrhenius constant

Ea = Activation energy

t = time /s

R= Gas constant 8.31

T= temperature /Kelvin

**k = Ae-Ea/RT**

If we take logs of both sides we get

**ln k = ln A – Ea/RT**

We can now plot a graph ln k (y axis) against 1/T.

What will the gradient of the line be equal to? .......................................

Using the same reaction as the iodine clock experiment, this time we keep all concentrations constant and adjust the temperature.(see experiment 2 in the previous pack)

1. From shared class burettes run into a small conical flask

6.0 cm3 0.5 M potassium iodide

2.5 cm3 0.001 M sodium thiosulphate

2.5 cm3 starch

1. Into a test tube put 6 cm3 0.01M peroxodisulphate. Place both the conical flask and test-tube into a water bath so they reach the same temperature
2. Simultaneously tip the peroxodisulphate into a small conical flask and start the stop watch. Swirl to mix thoroughly then allow to stand, whilst recording the time taken for the blue black colour to appear.
3. Repeat the experiment with the same quantities of chemicals but changing the temperature, (remember the starch-iodine complex becomes unstable at high temperatures.)

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Experiment No | Temp  / oC | Temp  / K | 1/Temp  K-1 | Time  s | 1/time = rate, k  s-1 | lnk |
| 1 |  |  |  |  |  |  |
| 2 |  |  |  |  |  |  |
| 3 |  |  |  |  |  |  |
| 4 |  |  |  |  |  |  |
| 5 |  |  |  |  |  |  |
| 6 |  |  |  |  |  |  |

**Calculations**

Note: minus the negative gradient gives a positive value for the activation energy

Activation energies are usually + 50 kJ mol-1 approx.

(The only reaction with a negative Ea is NO + O2)

Value for Ea =

Draw an energy level diagram for this reaction

Using collision theory discuss the significance of Ea

What factors do we need to control in this experiment to ensure an accurate result?

**Questions on rates of reaction and the Arrhenius equation**

See p12-13 **ln k = ln A – EA / RT** If you use log10 then a factor of 2.3 is needed, otherwise use natural ln.



1. The rate constant was found to vary with temperature as follows:-

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| T (K) | 293 | 313 | 333 | 353 |
| k | 0.0030 | 0.0216 | 0.122 | 0.567 |

Find the activation energy for this reaction.

1. Using the data from the following table plot a graph of rate of reaction vs. concentration. Use this to determine the order of reaction with respect to X and find the rate constant.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| [X] mol dm-3 | 0.100 | 0.210 | 0.285 | 0.420 | 0.540 | 0.700 |
| Rate | 0.0055 | 0.0116 | 0.0157 | 0.0231 | 0.0297 | 0.0385 |

1. The rate constant was found to vary with temperature as follows:-

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| T (K) | 290 | 310 | 330 | 350 |
| k x 103 | 1.01 | 5.04 | 20.7 | 72.2 |

Find the activation energy for this reaction.

**4)** Using the data from the following table, find the order of reaction with respect to A and B, the overall order of reaction and give the rate equation. Calculate a value for the rate constant.

|  |  |  |  |
| --- | --- | --- | --- |
|  | Experiment 1 | Experiment 2 | Experiment 3 |
| [A] mol dm-3 | 0.12 | 0.36 | 0.36 |
| [B] mol dm-3 | 0.04 | 0.04 | 0.20 |
| Rate of loss of A | 9.0 x 10-5 | 8.1 x 10-4 | 8.1 x 10-4 |

**Experiment : To find a suitable catalyst for the reaction CPAC 2c, 4a, 5a**

S2O82- + 2I-  2SO42- + I2

Oxidation half equation

Reduction half equation

Possible catalysts Cr (VI), Cr(III), Mn(VII), Mn(II), Fe(III), Fe(II)

Transition metal catalysts work because they can exist in a **variety of oxidation states**,

**Method**

1. From shared class burettes run into a small conical flask

2.0 cm3 potassium iodide

2.0 cm3 sodium thiosulphate

1.0 cm3 starch

1. Into a test tube put 4 cm3 of peroxodisulphate
2. Simultaneously tip the peroxodisulphate into a small conical flask and start the stop watch. Swirl to mix thoroughly then allow to stand, whilst recording the time taken for the blue black colour to appear.
3. Rinse out the conical flask very thoroughly. Why?
4. Repeat the experiment with the same quantities of chemicals but adding **1 drop** of catalyst solution in the lower or higher oxidation state to the conical flask prior to adding the peroxodisulphate.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  |  | Time taken (s) | Class average (s) | Did this act as a catalyst? |
| Control | |  |  |
| Chromium | Cr3+ |  |  |  |
| Cr2O72- / H+ |  |  |  |
| Manganese | Mn2+ |  |  |  |
| MnO4- / H+ |  |  |  |
| Iron | Fe2+ |  |  |  |
| Fe3+ |  |  |  |

Which of the three transition metals best catalysed the reaction……………

A possible mechanism involves **two redox reactions**:-

Write out the equation of Fe(III) with iodide:-

Write out the equation of Fe(II) with peroxodisulphate:-

Explain the mechanism. ………………………………………………………………………

………………………………………………………………………………………………...

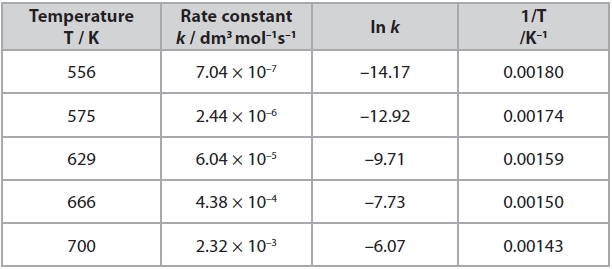
Which of the metals did NOT function as a catalyst at all……………………………………

When another class did the same experiment a week later the pattern was the same but the recorded times were different, suggest an explanation

………………………………………………………………………………………………………………………………………………………………………………………………………………………………………………………………………………………………………

**Questions**

**Q1.**The rate constant for this reaction was determined at various temperatures.   
The data from these experiments are shown in the table.

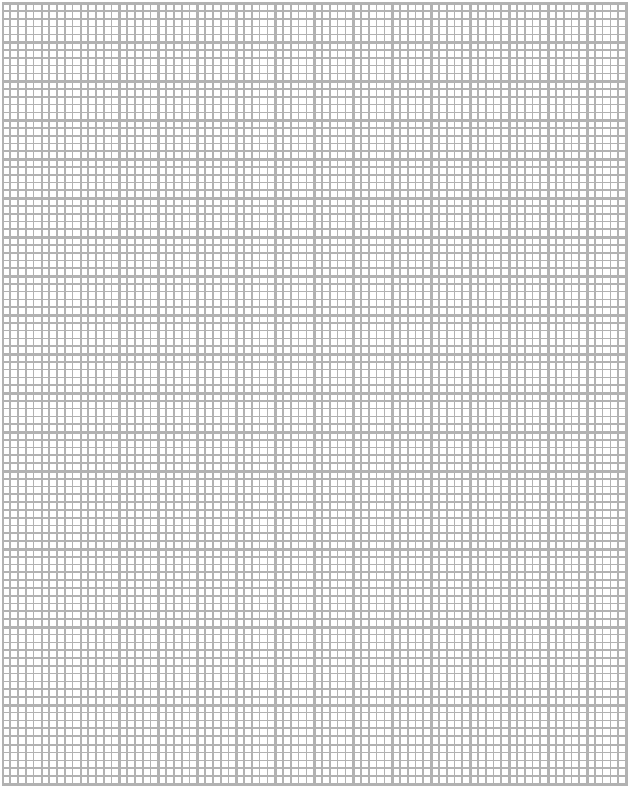


Use the data in the table to plot a graph and use it to calculate the activation energy, *E*a, of the reaction, giving your answer to an appropriate number of significant figures and stating the units.

The rate constant of a reaction, *k*, is related to the temperature, **T**, by the expression



**(5)**



Activation energy = ...........................................................

**Q2**

This question is about the decomposition of hydrogen iodide.



The decomposition of hydrogen iodide is catalysed by the heterogeneous catalyst platinum.

(a)  State the meaning of the term heterogeneous.

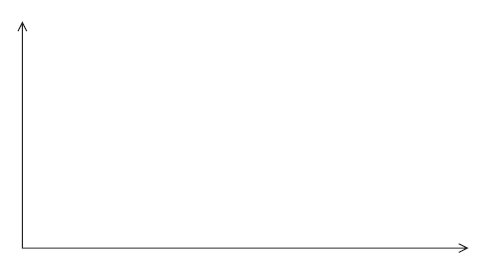
**(1)**

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(b)  (i)  Label the axes and draw a curve to show the Maxwell-Boltzmann distribution of molecular energies in a gas. Mark on your graph a suitable value for the activation energy, *E*a , for the reaction.

**(3)**



(ii)  Use the Maxwell-Boltzmann distribution to explain why a catalyst increases the rate of a reaction.

**(3)**

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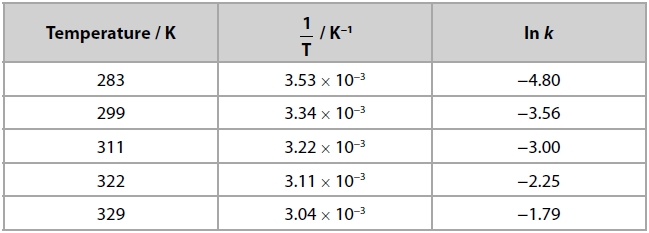
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**Q3.**This question is about the effect of changes in temperature on reactions.

An experiment to determine the activation energy for the reaction between magnesium and hydrochloric acid was carried out. The time taken for 0.100 g of magnesium to react completely when added to 20.0 cm3 of 1.00 mol dm−3 hydrochloric acid was recorded at different temperatures. A summary of the processed data is shown.

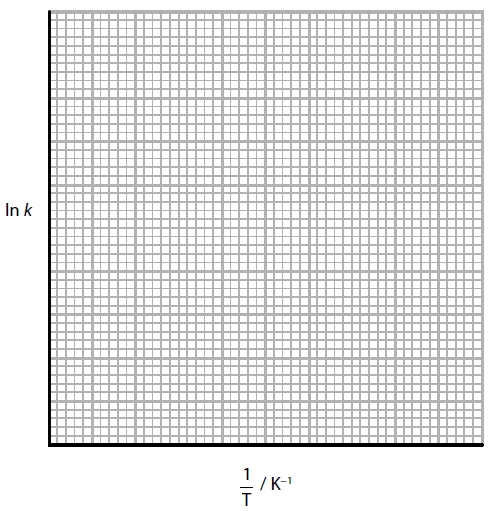


The activation energy, *E*a, of the reaction can be found using the equation:



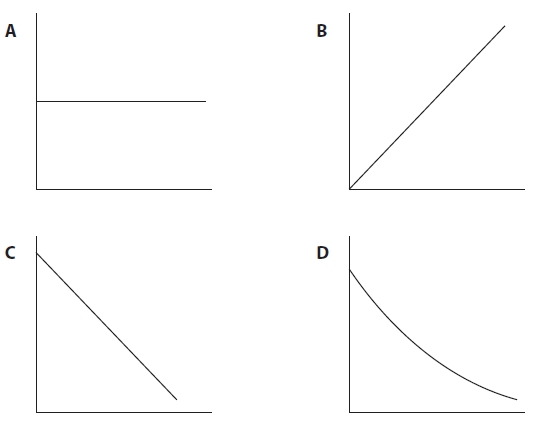
Use this data to plot a graph of ln *k* against  and hence determine the activation energy in kJ mol−1.

**(4)**

Activation energy = ........................................................... kJ mol−1

**Q4.**

Four sketch graphs are shown below.



(a)  Which could be a graph of rate of reaction, on the vertical axis, against the concentration of a reactant for a zero order reaction?

**(1)**

   **A**

   **B**

   **C**

   **D**

(b)  Which could be a graph of rate of reaction, on the vertical axis, against the square of the concentration of a reactant for a second order reaction?

**(1)**

   **A**

   **B**

   **C**

   **D**

(c)  Which could be a graph of the concentration of a reactant, on the vertical axis, against time for a first order reaction?

**(1)**

   **A**

   **B**

   **C**

   **D**

(d)  Which could be a graph of ln(rate), on the vertical axis, against reciprocal of temperature, 1/T, for a reaction?

You may use the equation    ln(rate) =  + constant

**(1)**

   **A**

   **B**

   **C**

   **D**

**(Total for question = 4 marks)**

**Q5.** The kinetics of the fast reaction below were investigated in a series of experiments.



(a)  Outline a titrimetric method that could be used to measure the change in concentration of compound **A** with time. Compound **A** is an alkali, whereas compounds **B**, **C** and **D** are neutral.

**(3)**

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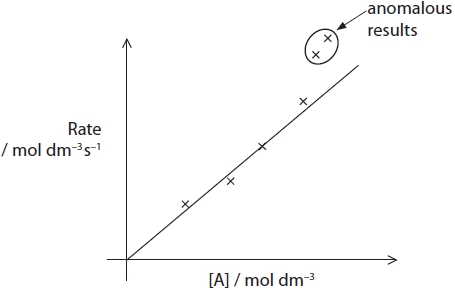
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(b)  The rate of the reaction was measured at several different concentrations of **A**, in the presence of a large excess of compound **B** and a constant amount of catalyst **X**, to find the order of reaction with respect to **A**. The results are shown on the graph below.



(i)  Explain how the graph confirms that the reaction is first order with respect to **A**.

**(1)**

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(ii)  Suggest an explanation, other than human error, for the two anomalous results circled on the graph.

**(3)**

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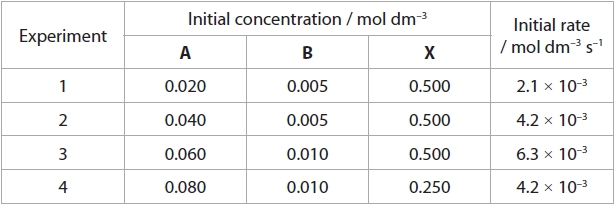
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(c)  In a second series of experiments, further data were collected using an initial rates method. These results are summarised in the table below.



(i)  Give **one** reason why obtaining these further data may be considered useful.

**(1)**

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(ii)  Deduce the rate equation for this reaction, explaining how you arrived at your answer.

**(5)**

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(iii)  Use your answer from (c)(ii), and appropriate data from **Experiment 4**, to calculate the value of the rate constant, *k*. Include units in your answer.

**(2)**

**(Total for question = 15 marks)**

**NAME ...........................……... HOMEWORK DEADLINE .....................**

Student Number ………… Chemistry Class ………

Student targets from **previous pack**

Chemical Kinetics II

|  |  |
| --- | --- |
| **Task** | Mark |
| **Notes** | /10 |
| Revision Notes | /10 |
| Exam questions at end of pack | /35 |
| Overall Grade for this work | A\* A B C D E U |

Student comments

Tutor comments

Tutor signature Date

Student targets for **next pack**