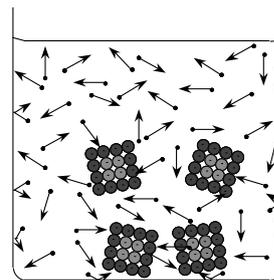
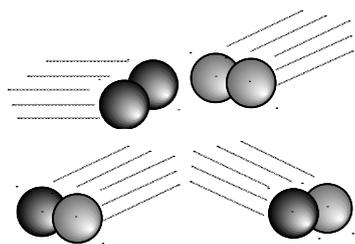


Higher Chemistry



Topic 2:

Reaction Rates

Student:

Lesson	Activities	Done	Checked
3.1 Collision Theory	1. Concentration and Collisions		
	2. Particle Size and Collisions		
	3. Temperature and Collisions		
	<i>Check Test</i> Score: / 8		
	<i>Home Practice</i> Score: / 10		
3.2 Catalysts At Work	1. Catalytic Converters		
	2. Heterogeneous Catalysts		
	3. Homogeneous Catalysts		
	<i>Check Test</i> Score: / 8		
	<i>Home Practice</i> Score: / 10		
3.3 Concentration And Rate	1. Considering the Investigation		
	2. The Experimental Work		
	3. Processing the Results		
	<i>Check Test</i> Score: / 8		
	<i>Home Practice</i> Score: / 10		
3.4 Temperature And Rate	1. Considering the Investigation		
	2. The Experimental Work		
	3. Processing the Results		
	<i>Check Test</i> Score: / 8		
	<i>Home Practice</i> Score: / 10		
3.5 Reaction Progress	1. Progress of a Reaction		
	2. Calculating the Rate		
	3. Comparing Reaction Progress		
	<i>Check Test</i> Score: / 8		
	<i>Home Practice</i> Score: / 10		
3.6 Activation Energy	1. Energy During Collision		
	2. Potential Energy Diagrams		
	3. Catalytic Effect		
	<i>Check Test</i> Score: / 8		
	<i>Home Practice</i> Score: / 10		
3.7 Energy Distribution	1. Distribution Graphs		
	2. Concentration and Temperature		
	3. Catalytic Effect		
	<i>Check Test</i> Score: / 8		
	<i>Home Practice</i> Score: / 10		
Consolidation Work	Consolidation A	Score: / 10	
	Consolidation B	Score: / 10	
	Consolidation C	Score: / 10	
	Consolidation D	Score: / 10	
	End-of-Section Assessment	Score: _____ %	Grade: _____

2.1 Collision Theory

This first lesson looks at how the effect of **concentration**, **particle size** and **temperature** on reaction rates can be explained in terms of **collisions** between reactant particles.

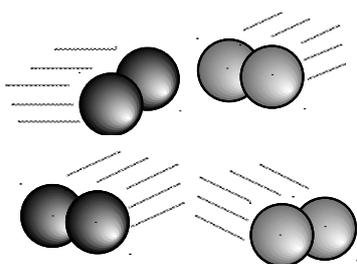
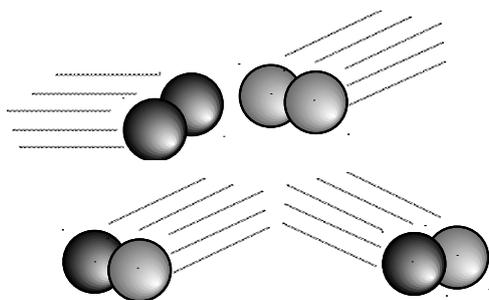
Concentration & Collisions

This activity looks at how the **collision theory** can explain why the concentration of reactants affects the rate of a chemical reaction.

The **Collision Theory** is simply common sense.

For a **reaction** to occur the particles of two reactants have to make contact.; they must **collide**.

The reactant particles must **collide** with enough **force** to break the existing **bonds** in their particles.



Gentle **collisions** just result in reactant particles bouncing apart **unchanged**.

Breaking bonds requires **energy**, and that energy comes from the force of the **collision**.

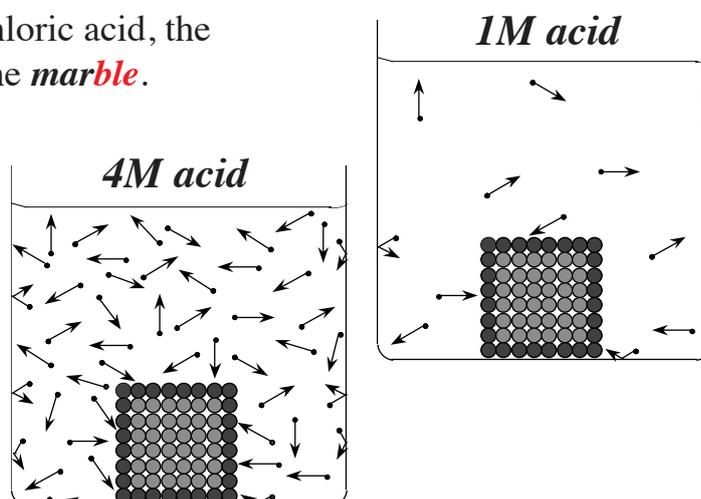
The Collision Theory states that substances can only react with each other if

- (i) their particles collide with each other, and
- (ii) the collisions have enough energy.

In the reaction between marble and hydrochloric acid, the **hydrogen ions** collide with the surface of the **marble**.

In **more concentrated** acid, there are more **hydrogen ions per cm³**, so there will be more **collisions per second**.

Not every **collision** will have enough **energy**, but there will be more **successful collisions per second** than before, so the **rate** of the reaction will be **increased**.



Particle Size & Collisions

This activity looks at how the **collision theory** can explain why the **particle size** of a **solid reactant** affects the rate of a chemical reaction.

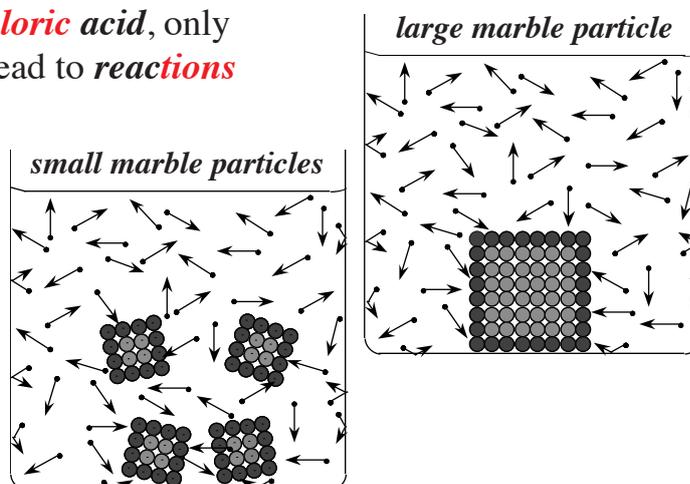
The word **particle**, in this context, refers to the form of the solid; lump, grains, powder etc

In the reaction between **marble** and **hydrochloric acid**, only **collisions** on the **surface** of the marble can lead to **reactions**

In the **large** marble particle, only **28** marble ions (CO_3^{2-}) are on the **surface**.

The mass of the **small** marble pieces is the same as the **large** marble, and so, therefore, are the **number** of marble ions,

This time, there are **48** marble ions on the **surface** of the small marble lumps.



With a **larger surface area** available, there will be **more collisions per second** which will result in a **faster reaction**.

Temperature & Collisions

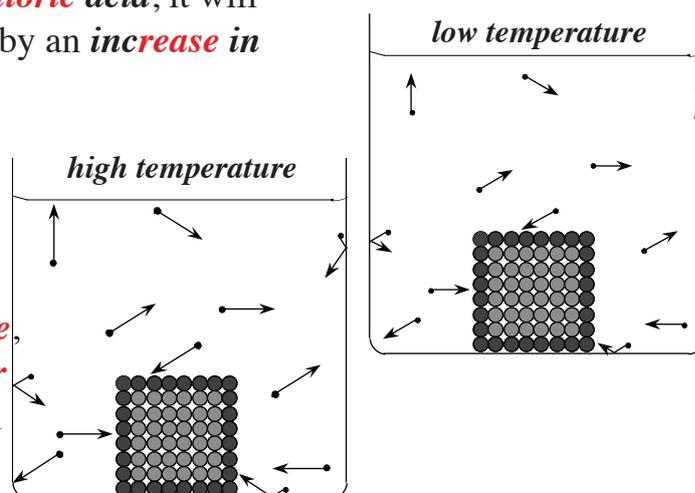
This activity looks at how the **collision theory** can explain why the **temperature** of the reactants affects the rate of a chemical reaction.

Increasing the temperature causes the reactant particles to move faster. This means that they **collide more often** and also that they collide with **greater kinetic energy**.

In the reaction between **marble** and **hydrochloric acid**, it will be the **hydrogen ions** that are most affected by an **increase in temperature**.

At the **higher temperature**, there will be **more collision per second** and this will result in a **faster reaction**.

More importantly, at the **higher temperature**, the **collisions** will be more 'violent', (**higher energy**), making a **reaction more likely** than with **collisions** at the **lower temperature**.



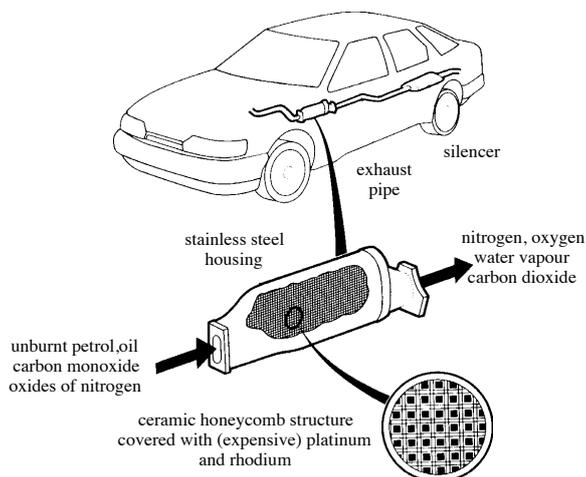
With **more collisions per second** taking place, and with a **larger proportion** of these **collisions** leading to a **reaction**, there will often be a **dramatic increase** in the reaction rate.

With many reactions, an increase of only 10°C is often enough to double the rate of the reaction.

2.2 Catalysts at Work

This lesson considers catalysts at work in different kinds of chemical reactions.

Catalytic Converters



This activity looks at how the **catalytic converter** in a car exhaust system converts polluting gases into non-polluting substances.

Converters are needed to change **harmful gases** into **less harmful gases**. Most have two sections:-

Section 1 - platinum & rhodium catalysts



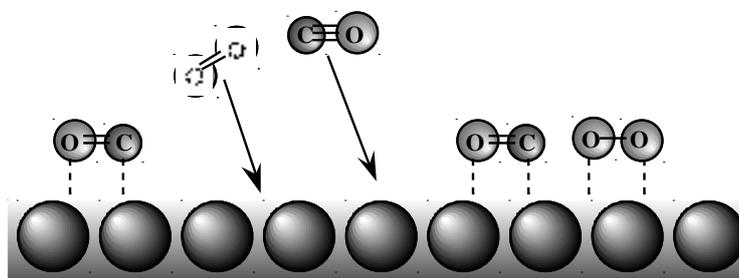
Section 1 - platinum & palladium catalysts



The gas molecules **collide** with the **catalyst** and stick to its **surface**. This is called **adsorption**.

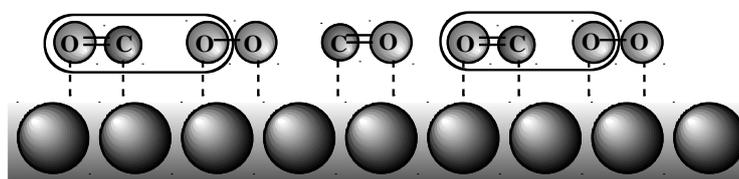
Weak bonds form between the **reactant** molecules and the **catalyst**.

This weakens the original bonds **within** the molecules.



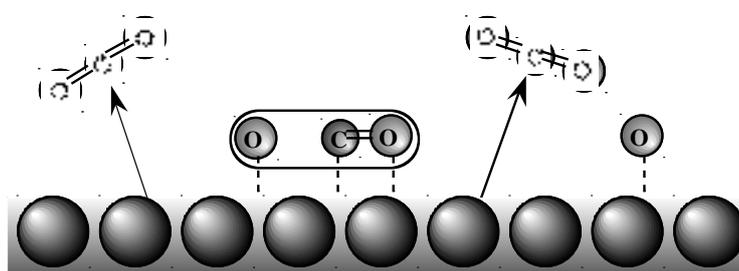
This makes it much 'easier' for the **reactant** molecules to react to form **products**.

'Easier' means the **collisions** can be **less violent**, so a reaction can occur at a **lower temperature** than normal, or will proceed at a **faster rate** than normal for that temperature.



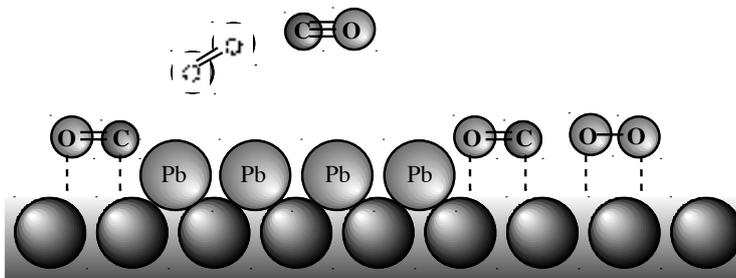
Once the **product** has formed, the **weak bonds** with the catalyst break and the product molecules depart.

The **surface** of the catalyst is then available to **adsorb** more reactant molecules and repeat the process.



Cars with **catalytic converters** must be able to run on **unleaded petrol**.

The **lead** atoms can form **strong bonds** on the **surface** of the **catalyst** and so prevent the **adsorption** of **reactant** molecules.



This effect is known as **poisoning** the catalyst. Catalysts of this type are often referred to as **surface catalysts** and benefit from a **large surface area**, so are often used as **powders** or sprayed onto a honeycomb structure. (This can also make a small amount of, very expensive, **transition metal** go much further).

Heterogeneous Catalysis

This activity is about examples of heterogeneous catalysis.

In heterogeneous catalysis the catalyst is in a different state from the reactants.

EXAMPLES OF HETEROGENEOUS CATALYSIS

<i>Process</i>	<i>Reactants</i>	<i>Catalyst</i>	<i>Important Product</i>

During the **industrial cracking** of large oil molecules the **powdered aluminosilicate** catalyst becomes coated with **carbon** (soot). This is another example of **catalyst poisoning**. It cannot be avoided, but the powder is continually **removed** and the soot **burned** off to allow the **catalyst** to be used again. This is known as **Regeneration**.

Homogeneous Catalysis

This activity is about examples of homogeneous catalysis.

*In homogeneous catalysis the catalyst is in the **same** state as the reactants, (usually that of **solution**).*

In ordinary life, many **reactions** take place in the **solutions** inside your body **cells** and are catalysed by **natural catalysts** called **enzymes**.

Enzymes are biological catalysts. They are complex protein molecules which speed up specific chemical changes by homogeneous catalysts.

ENZYMES AT WORK		
Industry	Enzyme or its Source	Change Achieved
	<i>invertase</i>	<i>sugar</i> → +
<i>wine making</i>		→
<i>yoghurt making</i>		<i>milk</i> →
	<i>yeast</i>	→ <i>beer</i>

*Hydrogen peroxide **solution** and Rochelle salt **solution** react together to produce **bubbles** of **oxygen** gas. The reaction is **extremely slow** at **room temperature**, but at 50°C the bubbles of oxygen are noticeable.*

***Pink** cobalt chloride **solution** ($\text{Co}^{2+}_{(\text{aq})}$), when added to the mixture, begins to turn **green** ($\text{Co}^{3+}_{(\text{aq})}$). At the same time, the **speed of the reaction increases** rapidly.*

*As the reaction subsides, the **pink** colour of $\text{Co}^{2+}_{(\text{aq})}$ returns. Like all true **catalysts**, the cobalt chloride is **unchanged at the end** of the reaction,*

2.3 Concentration and Rate

The purpose of this lesson is to investigate in more detail the way that the concentration of a reactant affects the rate of a chemical reaction.

* *State the aim of the experiment*

Procedure

* *Describe how the concentration of the potassium iodide solution was varied*

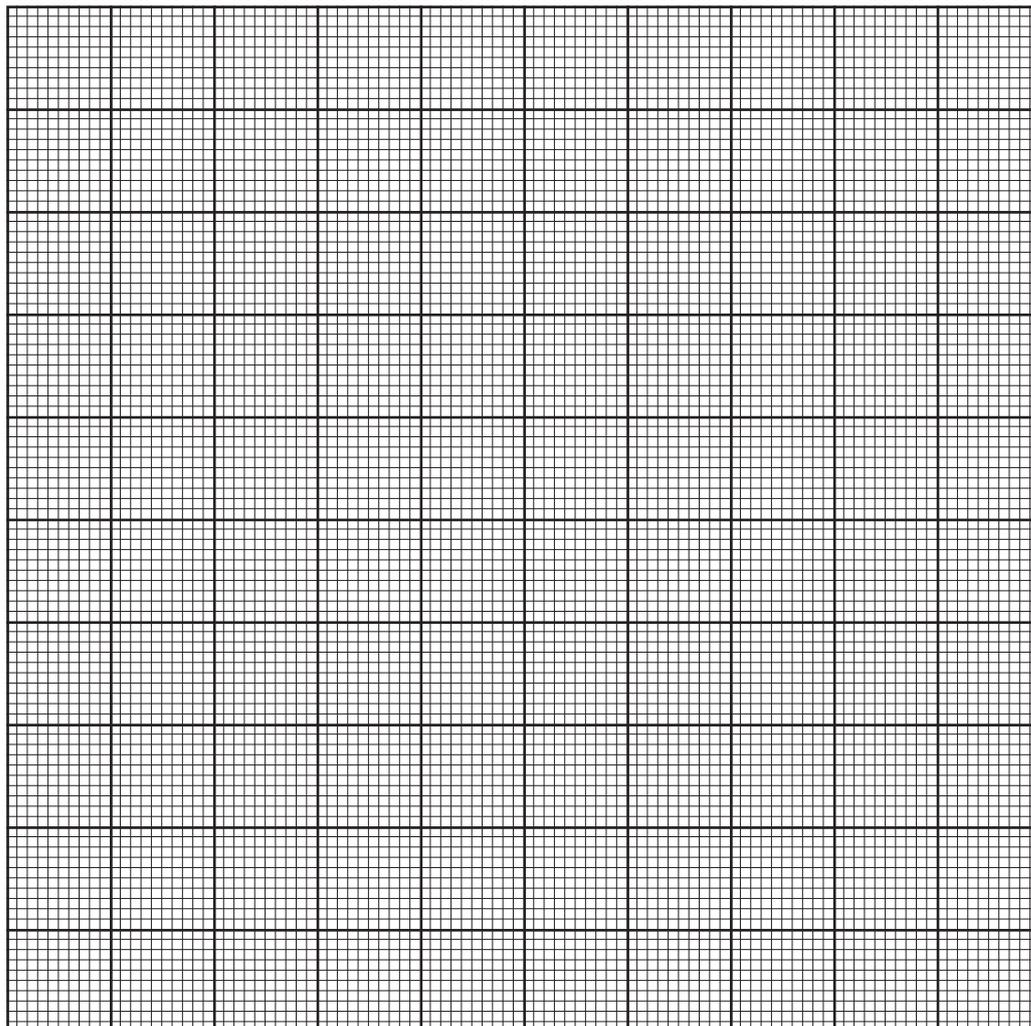
* *State two factors which had to be constant in the experiments.*

* *How was the rate of the reaction determined ?*

Results

* *Present Your results in a table:*

- * Draw a graph to illustrate your results



Conclusion

- * State the conclusion of the experiment

Evaluation

- * The procedure adopted to determine the reaction time was particularly effective and accurate. Suggest why this is the case.

- * Apart from allowing us to take the volume of potassium iodide as a measure of its concentration, what was the purpose of keeping the total volume of the reaction mixture the **same in each experiment**?

2.4 Temperature and Rate

The purpose of this lesson is to investigate in more detail the way that the temperature of the reactants affects the rate of a chemical reaction.

* *State the aim of the experiment*

Procedure

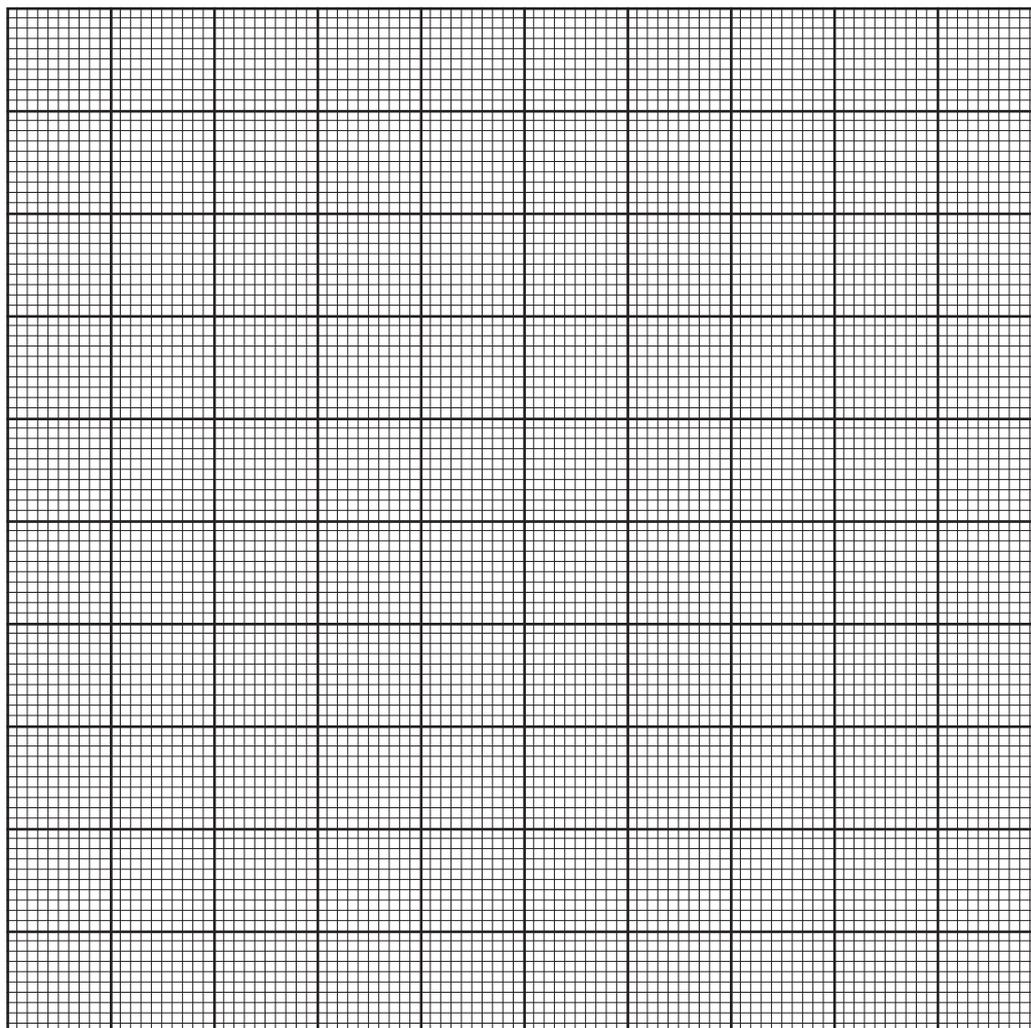
* *State two factors which had to be kept constant in the experiments.*

* *How was the rate of the reaction determined ?*

Results

* *Present your results in a tabular form:*

- * Draw a graph to illustrate your results



Conclusion

- * State the conclusion of the experiment

Evaluation

- * Why did the beaker in each experiment have to be dry?

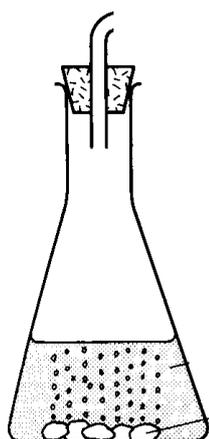
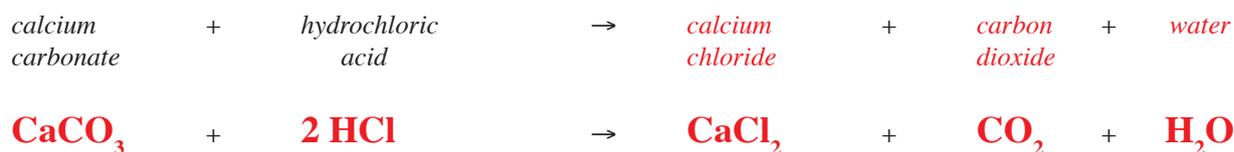
- * Suggest why it might be difficult to get an accurate reaction time if the experiment was carried out at room temperature

2.5 Reaction Progress

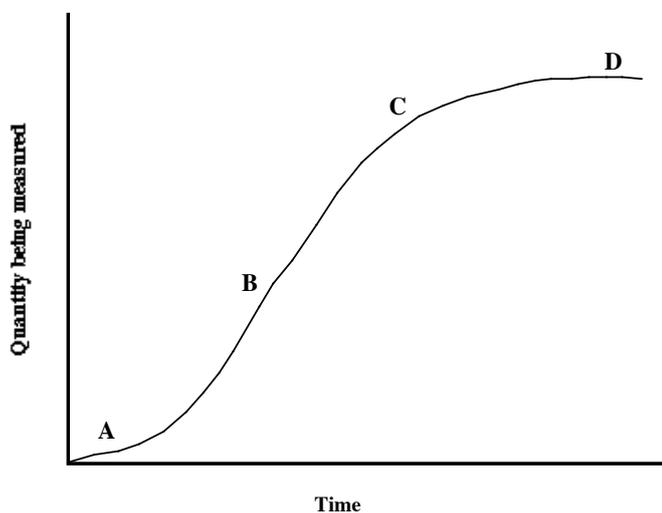
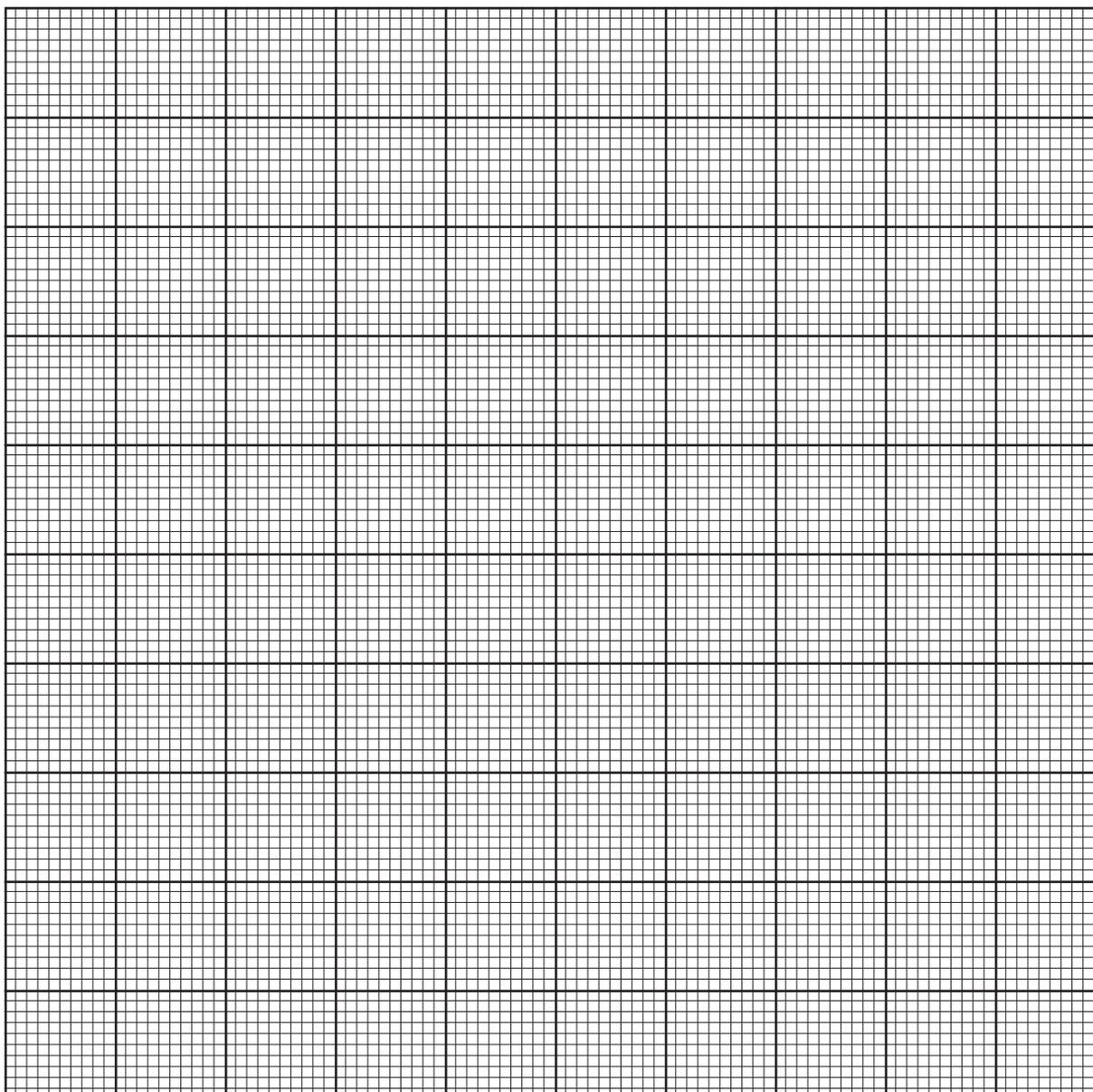
This lesson topic deals with some ways of following the progress of a chemical reaction.

Progress of a Reaction

The aim of the following experiment is to follow the progress of a reaction by recording the volume of gas produced at regular intervals.



<i>Time (min)</i>	
0	
2	
4	
6	
8	
10	



- A** *shallow slope* - many reactions are slow to get started
- B** *steep slope* - fast reaction rate
- C** *shallow slope* - reaction starts to slow down as chemicals are used up (their concentrations fall)
- D** *level slope* - reaction has stopped. One of the chemicals has been used up completely

Calculating the Rate

This activity examines how the rate of a reaction can be calculated from a progress graph.

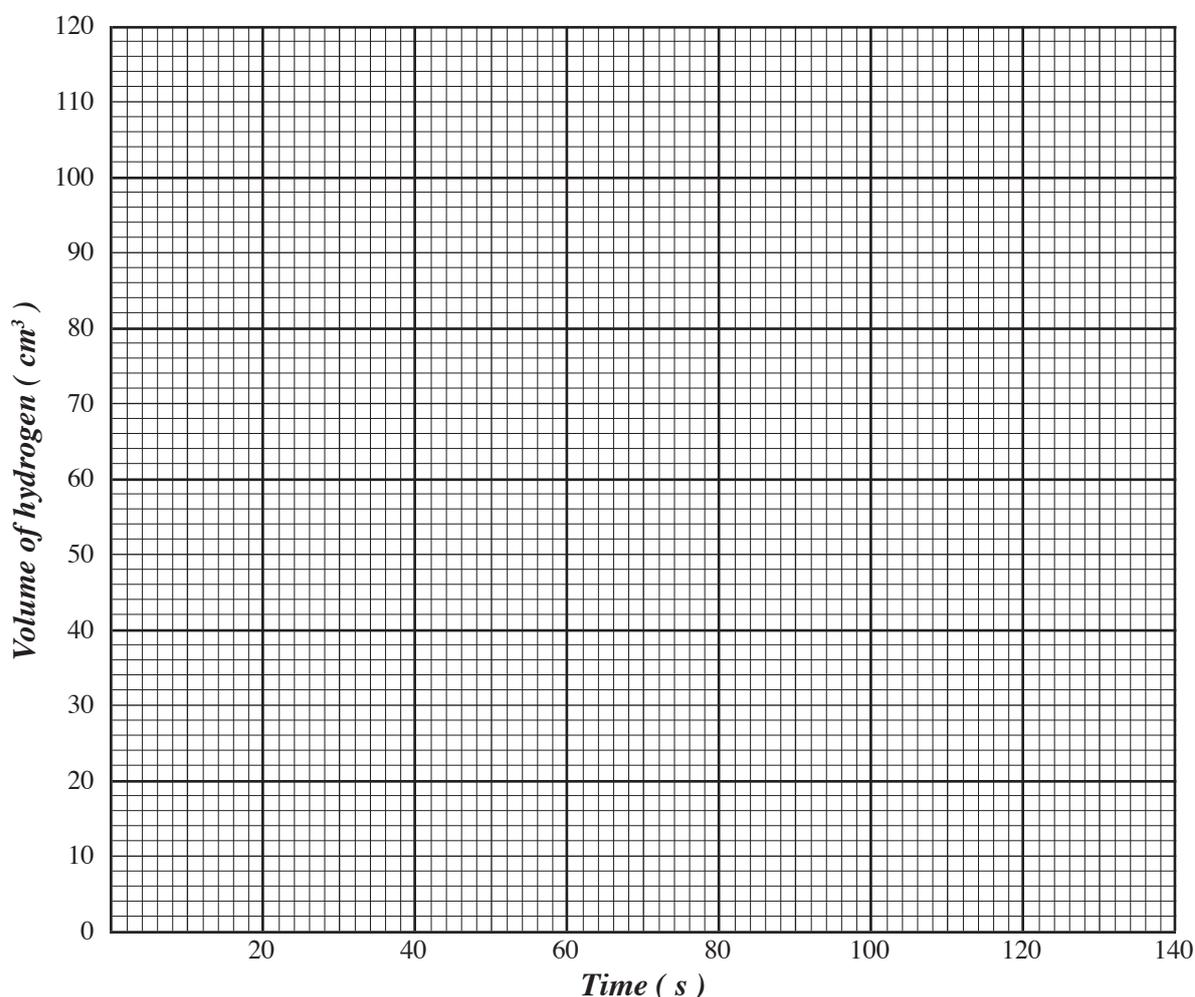
Rate of reaction is the **change in quantity** of a **reactant** or **product** per unit of **time**.

$$\text{average rate} = \frac{\text{change in quantity}}{\text{change in time}}$$

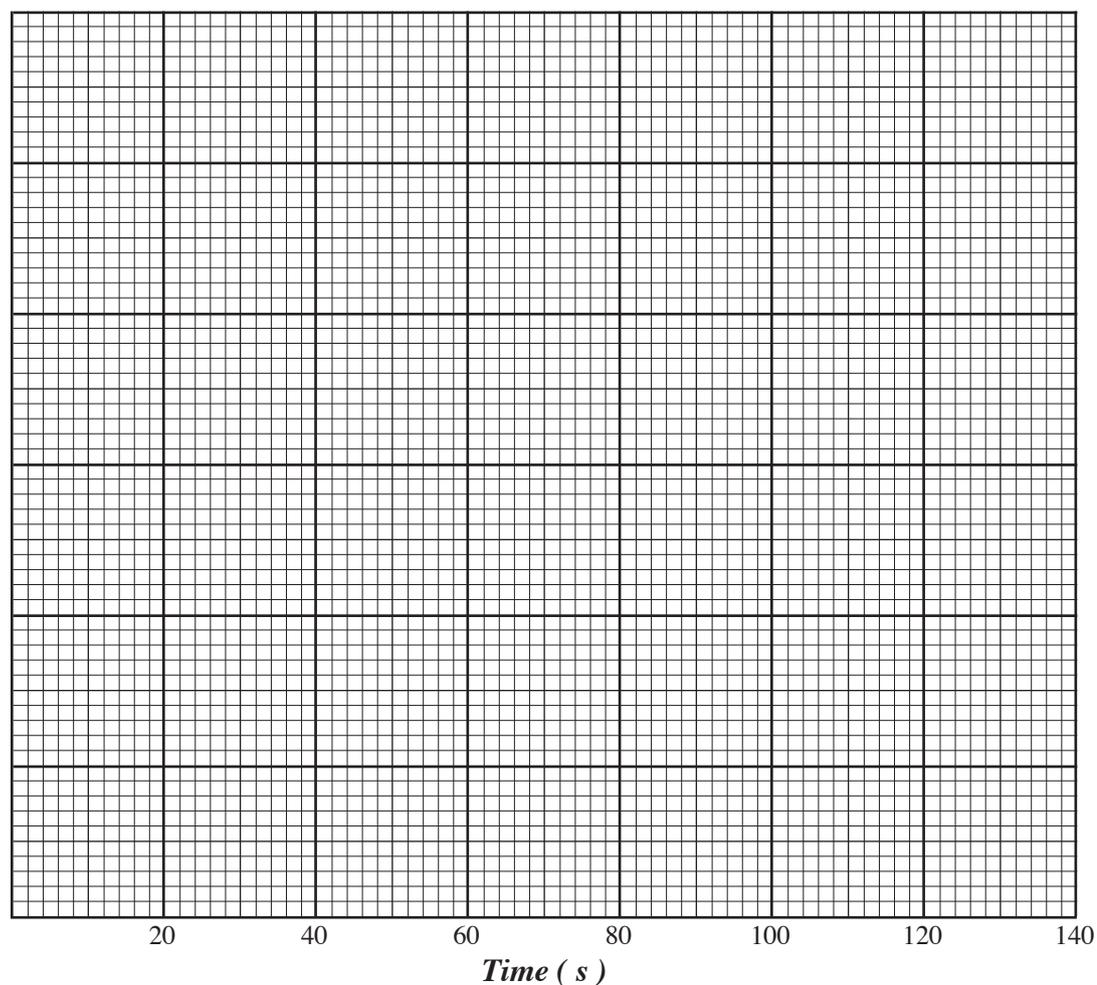
The **unit** used for **rate** depends on the **quantity** of the **reactant/product** that is being measured, and the **time scale** for the reaction.

e.g	weight loss (electrical balance)	grammes	g/s , g/min, g/hour
	gas volume (syringe)	ml or cm ³	cm ³ /s etc.
	concentration (colourimeter)	moles/litre	moles/l/s etc.

The reaction between **sulphuric acid** and **magnesium** produces **hydrogen** gas. The progress of the reaction can be monitored by **measuring** the **volume** of gas produced. The **Progress Graph**, below, can be used to **calculate** the rate of this reaction at different stages.



<i>Time interval</i> (s)	<i>Change in volume</i> (cm^3)	<i>Average rate</i> ($\text{cm}^3 \text{s}^{-1}$)
0 – 20		
20 – 40		
40 – 60		
60 – 80		
80 – 100		
100 – 120		
120 – 140		



The rate will be at a **maximum** near the **beginning** of the reaction, (when the **concentrations** of the **reactants** are at their **highest** level), will usually **drop** quite steadily (as the **reactant** concentrations **decrease**) and will eventually reach **zero** (once one of the reactants is used up completely.)

Comparing Reaction Progress

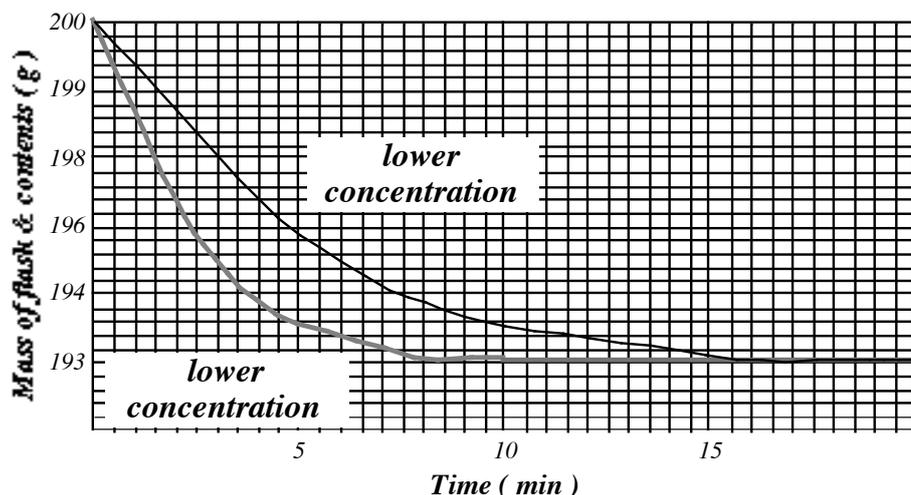
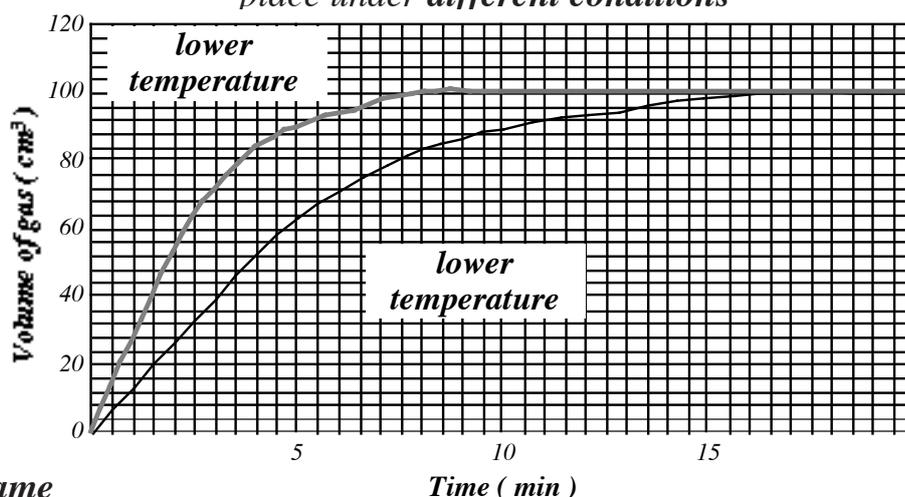
The purpose of this activity is to compare progress graphs for the **same reaction** taking place under **different conditions**

Example 1.

The **faster reaction** will produce **more gas** over the **same time** interval:- the slope will be **steeper**.

The **faster reaction** will **finish first**.

Both reactions have used the **same mass** of zinc, with the **same particle size**, with the **same volume** and **concentration** of hydrochloric acid. The **final volumes** of hydrogen gas will, therefore, be the **same**. The **faster reaction** was at a **higher/lower temperature**



Example 2.

The **faster reaction** will lose **more mass** over the **same time** interval:- the slope will be **steeper**.

The **faster reaction** will **finish first**.

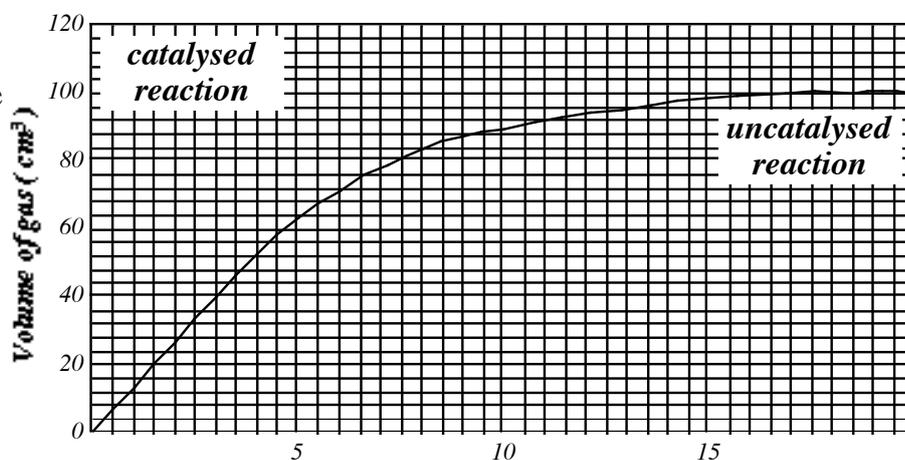
Both reactions have used the **same mass** of marble, with the **same particle size**, with the **same volume** of hydrochloric acid at the **same temperature**. The **final masses** of the flask & contents will, therefore, be the **same**. The **faster reaction** was at a **higher/lower concentration**

Example 3.

The **catalysed reaction** will be the **faster reaction** and will produce **more gas** over the **same time** interval:- the slope will be **steeper**.

The **catalysed reaction** will **finish first**.

Both reactions have used the **same mass** of zinc, with the **same particle size**, with the **same volume** and **concentration** of sulphuric acid at the **same temperature**, so the **final volume of gas** will be the **same**.



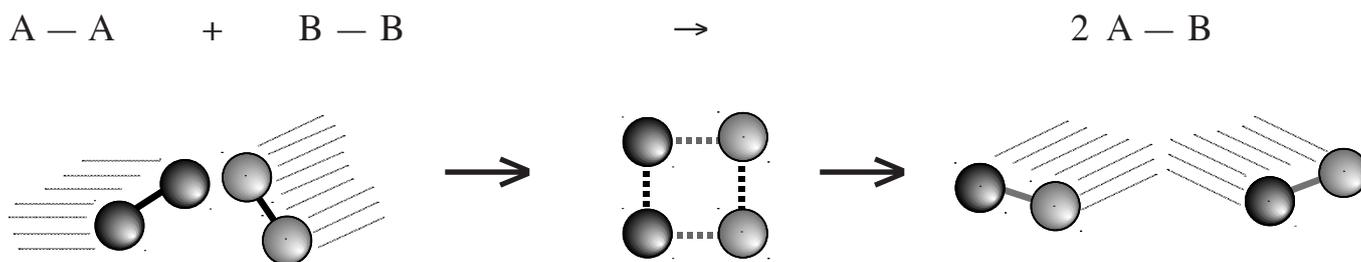
2.6 Activation Energy

This lesson explores the idea of activation energy and how it affects chemical reactions.

Energy During Collision

This activity looks at how the chemical energy of a reaction mixture changes as the reactant molecules collide and change to form the product molecules.

Imagine a very simple reaction between two *diatomic* molecules:



At the *start* of the reaction the *reactant* molecules have *kinetic energy* and *potential energy*.

The *potential energy* is mainly energy *stored* within their *bonds* and can also be called *chemical energy*.

As the *reactant* molecules begin to *collide*, strong forces of repulsion build up as their outer *electrons* repel each other.

The molecules begin to *slow* down, and their *kinetic energy* is converted into *potential energy*.

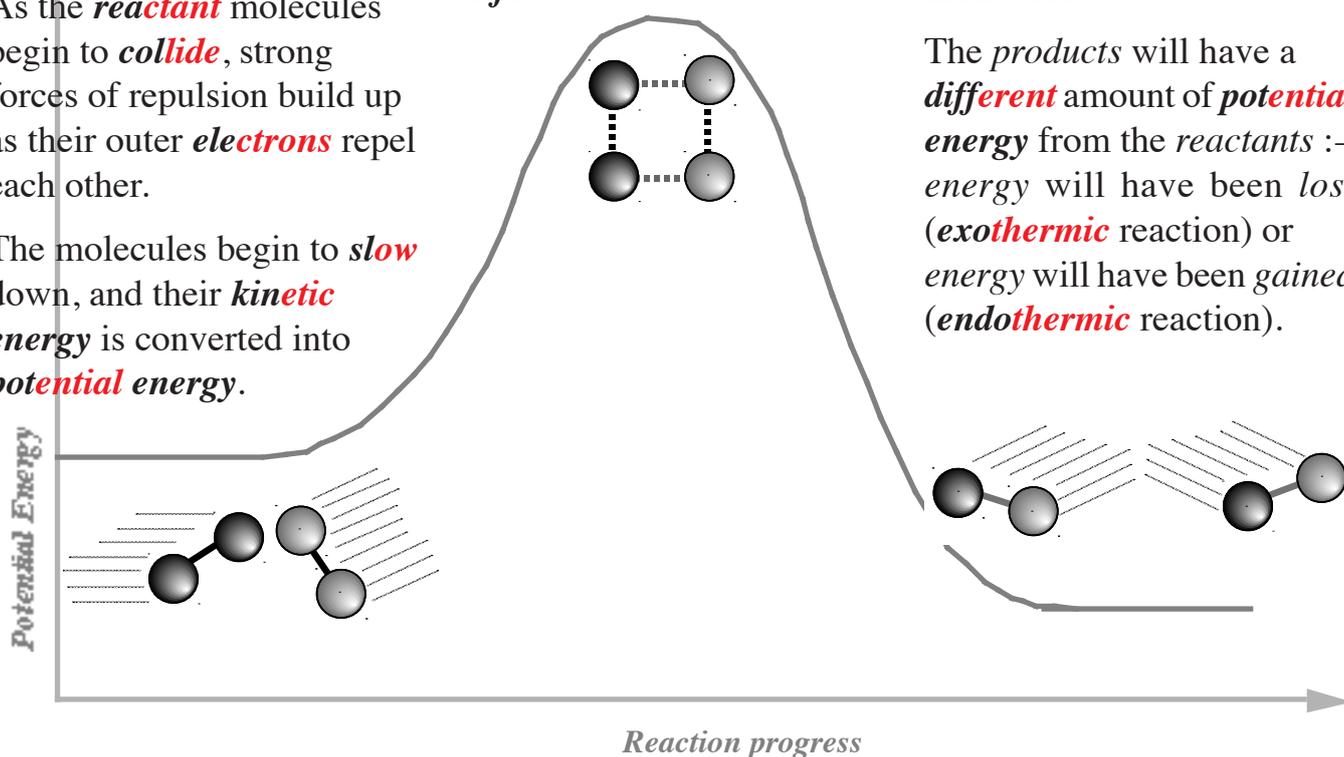
As the *molecules* crunch into each other and *stop*, their *kinetic energy* is now entirely *changed* to *potential energy*.

This *extra energy* can be enough to *break* existing *bonds* and allow *new bonds* to *form*.

If *new bonds* have *formed*, then as the *product* molecules spring apart *potential energy* will change back to *kinetic energy*.

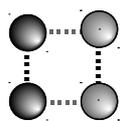
The *bonds* in the *product* molecules will be *different* from the original *reactant* molecules.

The *products* will have a *different* amount of *potential energy* from the *reactants* :- energy will have been *lost* (*exothermic* reaction) or energy will have been *gained* (*endothermic* reaction).

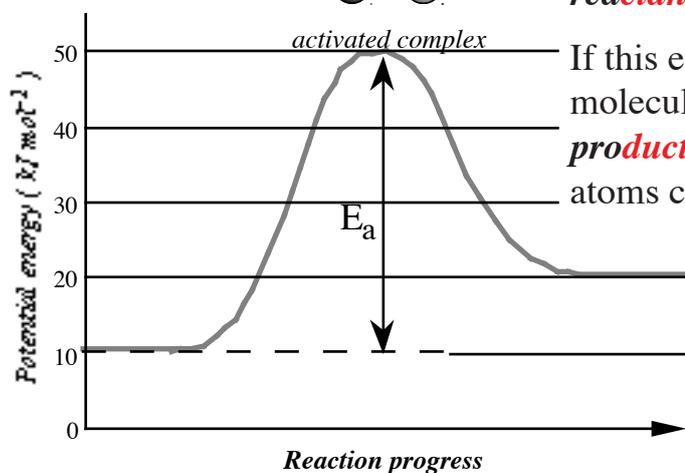


Potential Energy Diagrams

This activity takes a further look at the potential energy diagrams for chemical reactions



During **collisions**, **kinetic energy** is changed to **potential energy** and **stored** within the **bonds** of the **reactants**



If this energy is **large** enough, then the reactant molecules will **break apart** and go on to **form products**. In between, an **unstable arrangement** of atoms called the **activated complex** will exist.

If there is **not enough energy**, the reactant molecules will **separate** and proceed unchanged:- no reaction

In a sense there is an **energy barrier** which has to be overcome before reactant molecules can successfully

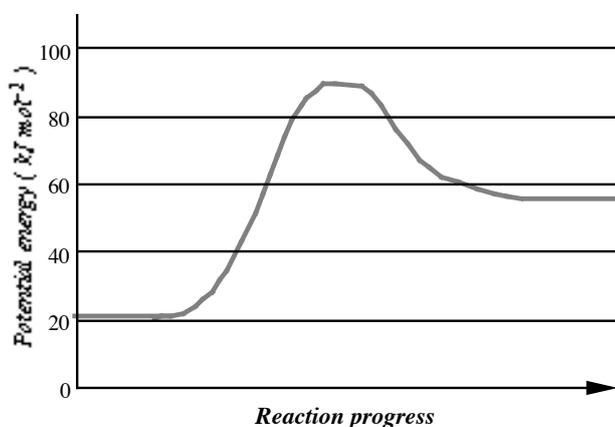
$$E_a = \quad \text{kJ mole}^{-1}$$

change into products. The **extra potential energy** that reactants need before they can react is called the **Activation Energy, E_a** .

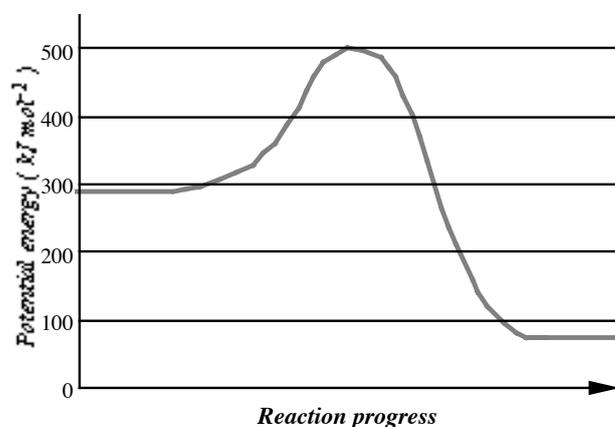
An activated complex is the unstable arrangement of atoms formed at the maximum of the potential energy barrier.

Activation energy is the additional potential energy which has to be achieved by colliding molecules to form an activated complex.

$$E_a = \text{P.E.}_{(\text{activated complex})} - \text{P.E.}_{(\text{reactants})}$$



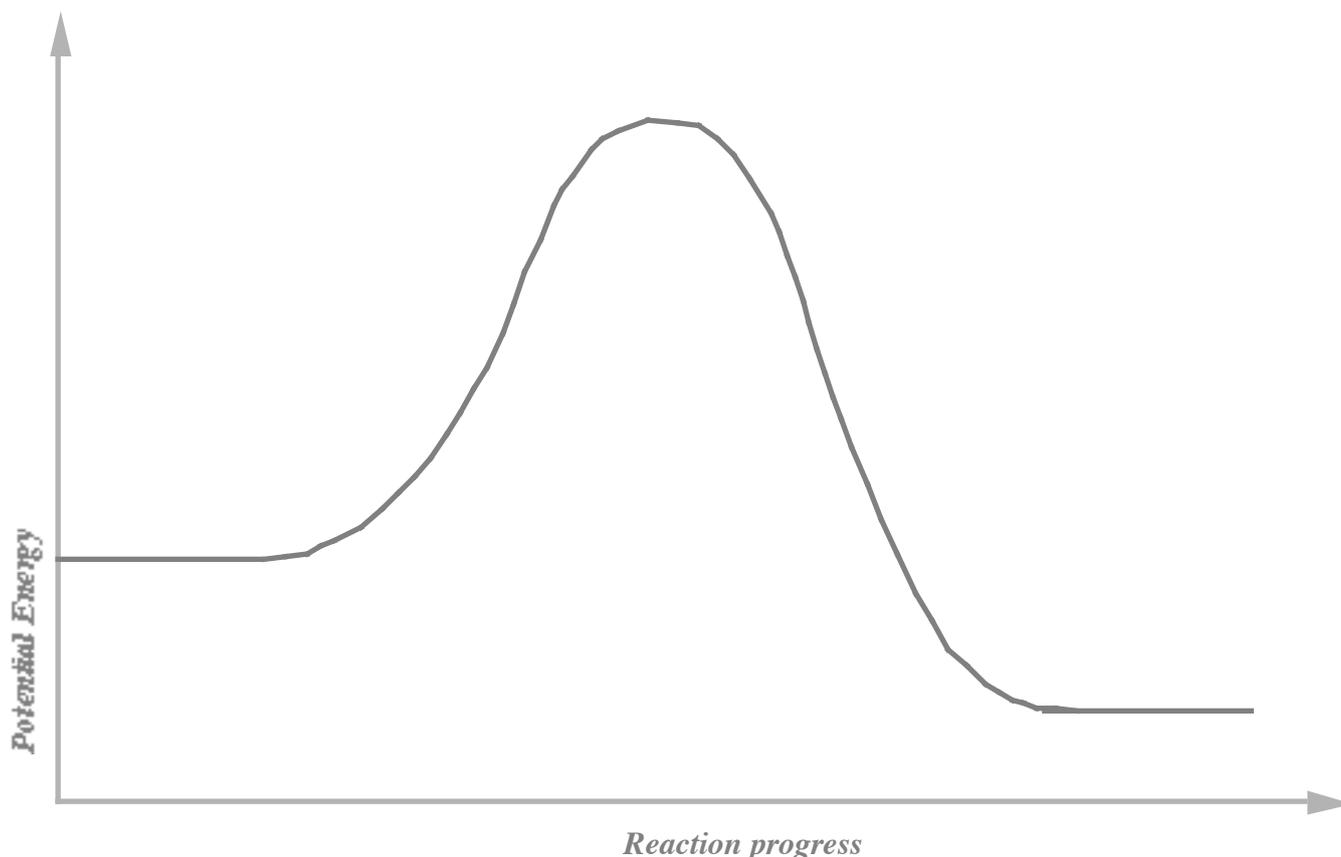
$$E_a = \quad 70 \quad \text{kJ mole}^{-1}$$



$$E_a = \quad 200 \quad \text{kJ mole}^{-1}$$

Potential Energy Diagrams

This activity explains the effect of a catalyst in terms of activation energy.



A **catalyst** provides a **different** reaction pathway:- a **different** series of **collisions** which will still change the **reactants** into **products**.

A **catalyst** lowers the **activation** energy :- the alternative route is 'easier', meaning that **less potential** energy is needed before **bonds** break.

Catalysed reactions can be done at **lower** temperatures :- the reactants need **less kinetic** energy to ensure a successful collision.

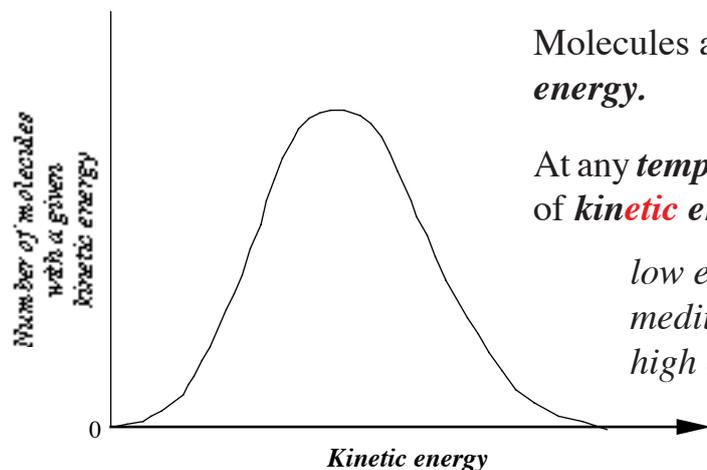
Catalysts speed up a reaction by providing an alternative reaction pathway with a lower activation energy

2.7 Energy Distribution

This lesson looks at the effect of a distribution of kinetic energies on the rate of a reaction

Distribution Graphs

This activity looks at distribution graphs for kinetic energies



Molecules are always **moving** - they have **kinetic energy**.

At any **temperature**, however, molecules can have a **range** of **kinetic energies**:-

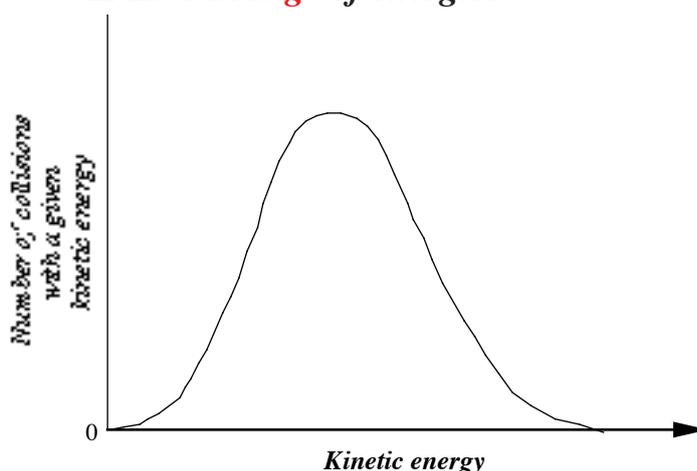
low energy	- very few
medium energy	- most
high energy	- very few

Similarly, **collisions** between molecules will have a **range of energies**

Somewhere along the x-axis there will be the value for the **minimum kinetic energy needed** for a **collision** to lead to a **successful reaction** - the **activation energy**.

All collisions with this energy, or above, will lead to a reaction. Shading this area on the graph gives an indication of the **number** of successful collisions.

The larger this number, the faster the reaction

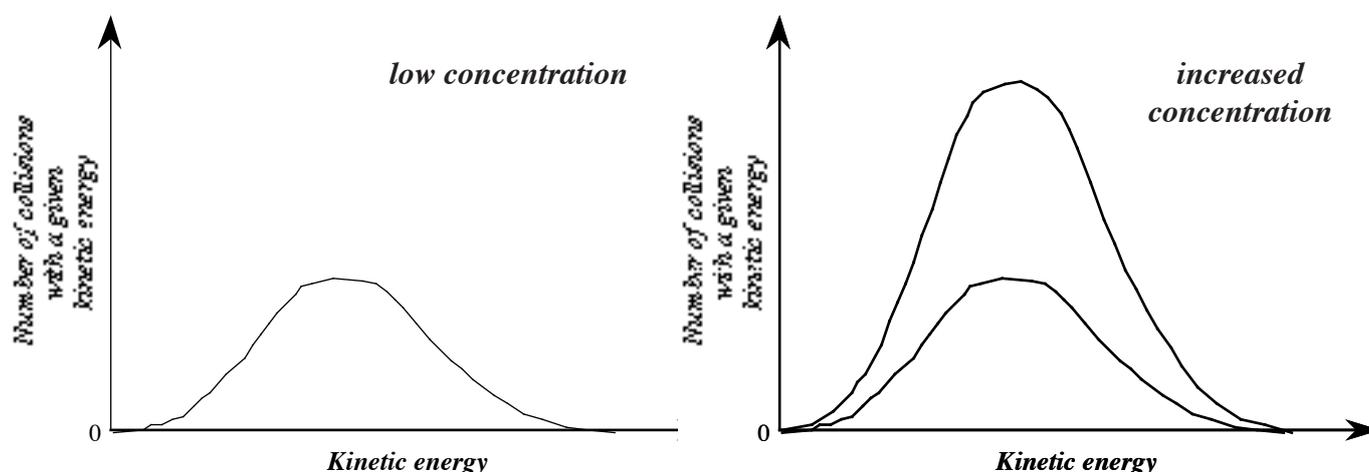


The normal method for **increasing the kinetic energy** of reacting particles is to **increase the temperature** - add **heat energy**. You should be aware, however, that many reactants are capable of taking in energy in other forms:-

- e.g. **light energy** - photosynthesis, **sunlight** provides the energy needed to speed up the reaction between CO_2 and H_2O .
- photography, light provides the energy to help convert silver ions back into silver atoms to produce the dark patches on a film negative
 - substitution, the reaction between **alkanes** and **bromine** requires sunlight
 - H_2 / Cl_2 the slow reaction between H_2 and Cl_2 can be speeded up (made **explosive**) using a UV camera flash

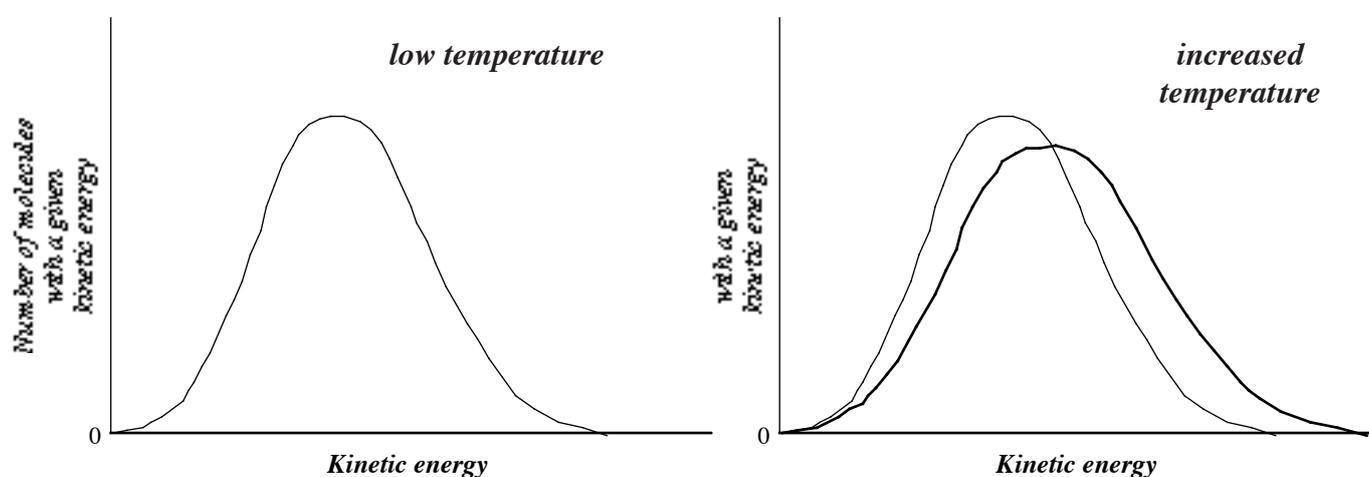
Concentration & Temperature

This activity looks at how increasing the concentration or increasing the temperature affects an energy distribution graph.



Increasing the concentration simply **increases the number of collisions** across all possible values of **kinetic energy**. For example, **doubling the concentration** **doubles the height** of the graph everywhere.

This, of course, **increases the number of collisions with activation energy** (or above) which **increases the number of successful collisions** and **increases the rate** of the reaction.



When you use a **thermometer** to measure the **temperature** of a substance you are really measuring the **average kinetic energy** of its particles. A **few** particles have **below average** energy, **most** have **close to average** while a **few** have **above average** energy.

Increasing the temperature causes all the molecules to move **faster**, to have **increased kinetic energy**.

The effect on the graph is to **spread** it out to **higher** values of **kinetic energy**. There will always be **a few** particles with **well below average energy** so the graph still starts at **zero**.

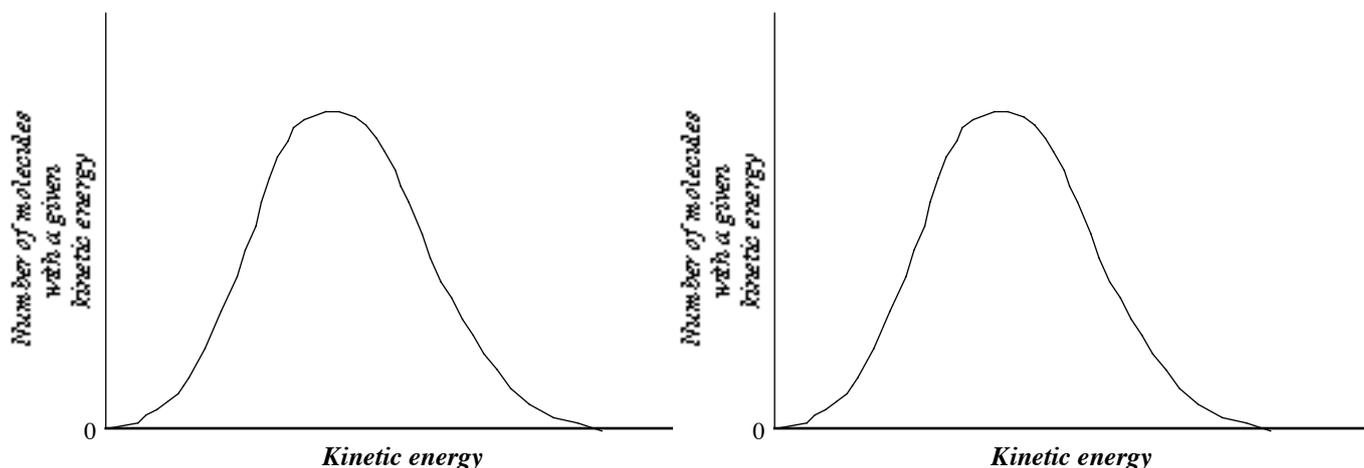
There will be **more** particles at the **high** end of the energy scale whilst there will have to be a similar **decrease** in particles in the middle. **The height of the graph will be lower**.

This, of course, *increases the number of collisions with **activation energy** (or above) which increases the number of **successful** collisions and increases the **rate** of the reaction.*

A small shift to the **right** can bring a *very large number* of particles over the **activation energy 'barrier'**. A *small increase in temperature* can often have a dramatic effect on the **rate of a reaction**. Often a 10°C rise in temperature is enough to **double** the rate of reaction.

Catalytic Effect

This activity explains how a catalyst speeds up a reaction without having any effect on the energy distribution of the molecules



Adding a **catalyst** has no effect on the **number of collisions** or on the **kinetic energy** of these collisions.

A **catalyst** provides an **alternative route** with a **lower activation energy**. Many **collisions** that, previously, had been lacking the **energy** needed to lead to a **reaction** will now prove to be **successful**.

This, of course, *increases the number of collisions with **activation energy** (or above) which increases the number of **successful** collisions and increases the **rate** of the reaction.*

UNIT 1. Chemical Change & Structure

Topic 2: Reaction Rates

Following the course of a reaction

1. Reactions can be followed by measuring changes in concentration, mass and volume of reactants or products.
2. The **average** rate of a reaction, or part of a reaction, can be **calculated** from initial and final **quantities** and the **time** interval
3. The rate of a reaction, or part of a reaction, is proportional to the **reciprocal** of the time taken.

Plot graphs of changes against time; slope = rate

Average rate = (change in quantity) / time

Rate = 1 / time

Factors affecting rate

4. The rates of reactions are affected by changes in **concentration**, **particle size** and **temperature**.

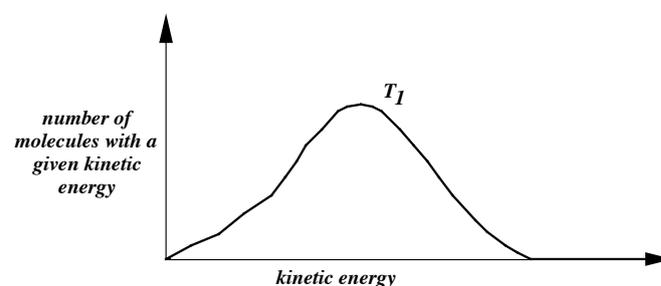
5. The **collision theory** can be used to **explain** the effects of concentration and particle size (surface area) on reaction rates

concentration ↑ number of collisions ↑
particle size ↓ surface area ↑ number of collisions ↑

6. Temperature is a measure of the average **kinetic energy** of the particles of a substance.

7. The **activation energy** is the minimum kinetic energy required by colliding particles before reaction will occur.

8. Energy distribution diagrams can be used to **explain** the effect of changing temperature on the kinetic energies of particles.



9. The effect of temperature on reaction rate can be explained in terms of an increase in the number of particles with energy greater than the activation energy.

10. With some chemical reactions **light** can be used to increase the number of particles with energy greater than the activation energy.

e.g. photosynthesis
photography

Catalysts

11. Catalysts can be classified as either **heterogeneous** or **homogeneous**.

heterogeneous - different state from reactants
homogeneous - same state as reactants

12. Catalysts are used in many industrial processes.

13. *Heterogeneous catalysts work by the **adsorption** of reactant molecules.*

Often called 'surface' catalysts

14. *The surface activity of a catalyst can be reduced by **poisoning**.*

Impurities adsorbed strongly onto surface.

15. *Impurities in reactants mean that industrial catalysts have to be **regenerated** or **renewed**.*

Made predictions.

16. ***Catalytic convertors** are fitted to cars to catalyse the conversion of poisonous carbon monoxide and oxides of nitrogen to carbon dioxide and nitrogen.*



17. *Cars with catalytic convertors only use "**lead-free**" petrol to prevent poisoning of the catalyst.*

18. ***Enzymes** catalyse the chemical reactions which take place in the living cells of plants and animals.*

19. *Enzymes are used in many industrial processes.*

Potential energy diagrams

20. *A potential energy diagram can be used to show the energy pathway for a reaction.*

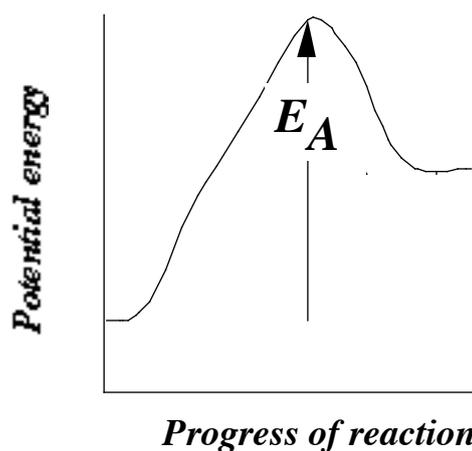
21. *The **activated complex** is an unstable arrangement of atoms formed at the maximum of the potential energy barrier, during a reaction.*

22. *The **activation energy** is the energy required by colliding molecules to form an activated complex.*

23. *The activation energy can be calculated from potential energy diagrams.*

24. *The effect of a catalyst can be **explained** in terms of alternative reaction pathways with lower activation energy.*

25. *A potential energy diagram can be used to show the effect of a catalyst on activation energy.*



$$E_a = \text{maximum energy} - \text{energy of reactants}$$