Energy

Teacher Notes

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Energy

Why focus on G&T and higher achievers?

Within the education system every child has the right to develop their learning so as to maximise their potential.

These exercises are designed to give students enthuse and enrich activities that although related to the curriculum are in fact taking the learning experience to the next level whilst also showing chemistry in a familiar context. This has been found to be a successful model for not only improving learning but also for raising levels of motivation. Higher achieving students can find the restraints of the standard curriculum to be demotivating leading to underachievement.

The different activities are designed to improve a number of skills including practical work/dexterity, thinking/analysis skills, literacy, research activities, use of models and teamwork. Students should also gain confidence through the activities and improve the ability to express themselves.

Some of the activities would appear to be complex for KS3, however at this stage in their learning high achieving students are open to new concepts and are ready to explore issues without pre-conceptions. They are keen to link ideas and develop concepts and understanding. It can prove to be an uplifting experience.

Introduction

Knowledge of energetics can help understand the process of a reaction and the conditions under which it operates. It also develops knowledge of the link between individual bond energies and the net energy gain or loss as well as the realisation that energy is not just heat.

This programme is designed to develop students understanding of energetics as well as developing thinking and research skills.

Topic	Type of activity	Summary	Timing (mins)	KS3	KS4	KS5	Page
Heating copper(II)sulfate	Practical	Experiencing the endothermic and exothermic properties of a reversible reaction.	30	√	√		6
Endothermic solid-solid reaction	Demonstration	A dramatic demonstration illustrating endothermic reactions.	10	√	√	√	15
Spontaneous exothermic reaction	Demonstration	A spontaneous reaction between a fuel and an oxidising agent.	10	√	√	√	17
Exothermic or endothermic	Practical	Measuring energy transfers in a range of reactions.	30	√	√		18
Measuring heat energy of fuels	Practical	An investigation looking at the varying amount of energy	60	√	√	√	24

		produced by the combustion of a range of alcohols. The practical results are then compared to calculated theoretical values.					
Chemiluminescence	Practical Demonstration	A simple experiment to illustrate that energy can be released in forms other than heat – in this case light.	10	√	√	√	28
Diagnostic test	Summative assessment	A test of knowledge and understanding.	20-40				31

The first activity looks at a reversible reaction and how the energy gain or loss is affected by the change in direction. It introduces the concept of exothermic and endothermic reactions with a colourful example. It also acts to illustrate water of crystallisation and at a higher level it can be used to develop the idea of ligands.

Activity 2 looks more carefully at the endothermic process using an example that can fall to -30^{o}C. The demonstration offers extension activities through discussion as to how two solids can react together to form a ‘liquid’(actually a solution/suspension) and a gas. It again brings in the idea of water of crystallisation as well as the rearrangement of substance upon reaction. There is even the opportunity to illustrate how to test for ammonia (alkaline gas and its reaction with concentrated hydrochloric acid). Exothermic reactions are explored in the next activity. The reaction between potassium manganate(VII) and glycerol is spontaneous as the activation energy is less than the ambient energy at room temperature and therefore it does not require energy to activate the reaction. It is also a good example of the reaction between an oxidising agent and a fuel as well as linking to the change in colour of the manganate with a change in oxidation state. It therefore links to many levels in the curriculum. Students are given the opportunity to explore, and practice the skill of observation, these properties in a class practical in activity 3.

Measuring the heat produced in reactions, in this case the combustion of alcohols as fuels, allows for the development of investigative skills. The practical can be completed on many levels with various levels of help. There is also a calculation component to the practical that opens up a wide area of discussion and evaluation. The calculation focuses on the theoretical value for the energy of the reaction and then why it differs from the ‘real’ value.

It is important that students realise that energy can be absorbed or released in different forms such as light and pressure as well as heat. It is always worth a discussion about the heat, light and sound produced by an exploding hydrogen/oxygen balloon. The chemiluminescent example helps to reinforce this fact.

The diagnostic test acts as a summative test to explore the level and knowledge of the topic.

Aims and objectives

The aims and objectives of these activities are:

- Developing questioning skills through problem solving.
- Exploring the use of models to expand understanding

- To appreciate that all models have limitations dependent upon the knowledge available and the application.
- Develop practical skills and dexterity.
- Promote independent learning and research skills.
- Chemistry topics:
    - Collision theory.
    - Energy of reactions
    - Activation energy
    - Exothermic reactions
    - Endothermic reactions
    - Energy from fuels
    - Energy released in other forms than heat

These exercises can be used with key stages 3 and 4, and to some degree key stage 5, as indicated on the *Possible Routes*.

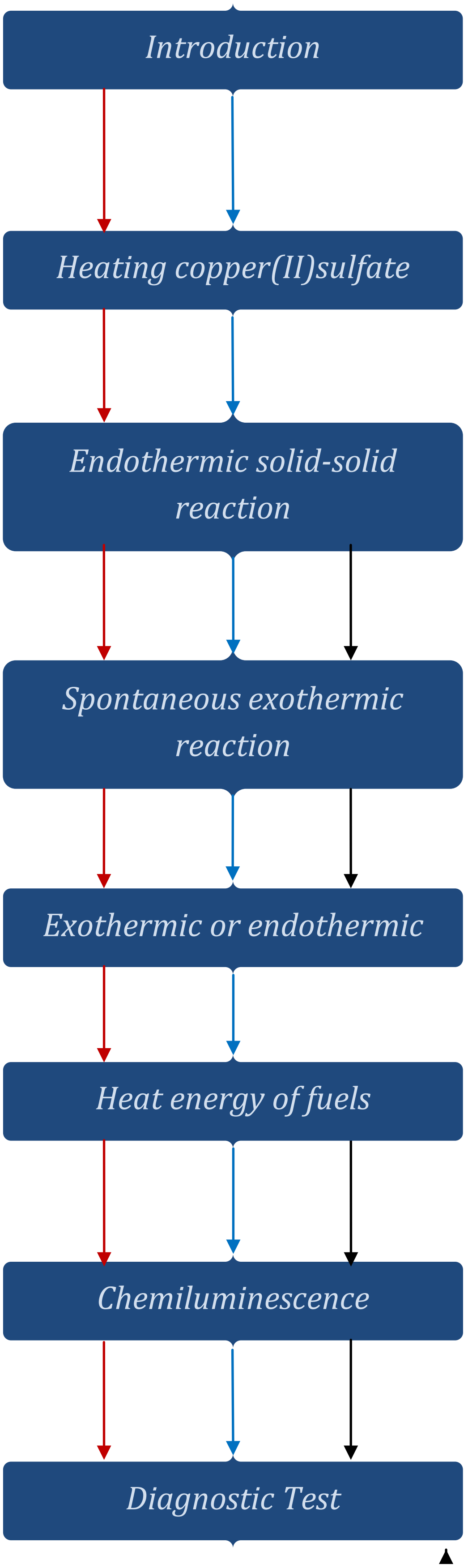
These activities have proved very successful with key stage 3 students who have followed the prescribed pathway and have been stimulated into further independent learning.

At key stage 4 it has enhanced understanding of energetic and developed skills in calculation of reaction energies, as well as reinforcing ideas about bonding.

This exercise provides a reinforcement and revision tool for a number of topics from the A level syllabus.

At all levels there is promotion of questioning skills, independent learning and research skills.

Possible routes



The introduction leads into a practical experiment that explore a reversible reaction with an exothermic and endothermic component

Moving forward to further develop the ideas around endothermic reaction with a dramatic demonstration with possibilities of extension

This activity seeks to explore the ideas of exothermic reactions and allows for extension

Having developed the theory this activity allows the students to develop their skills of observation over several reactions

Heat energy of fuels is a broad ranging investigation that develops the research and evaluation skills of students. It also introduces quantitative calculations

Chemiluminescence is used to illustrate that the energy involved in reactions does not solely involve heat

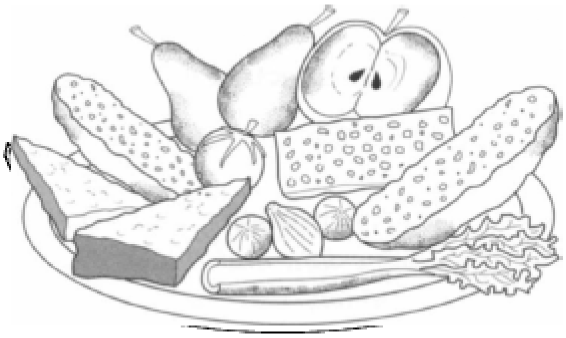
The diagnostic test looks to assess the level of knowledge and understanding of students.

Key

→	KS3 G&T
→	KS4
→	KS5
→	KS4 G&T

Introduction to Energy Changes

You are a walking, talking chemical reaction. Every step you take, every thought and breath, involves a series of physiological changes that makes use of chemical reactions - and you don't even have to think about it! The energy needed to ensure that these reactions occur comes from the body creating it from the food that we eat.



Food provides energy for the chemical reactions

Fats and carbohydrates are important biological fuels - the chemical energy stored in these foods can be converted into heat and movement energy, electrical energy in brain and muscles, and of course sound energy. Protein foods are essential for tissue growth and cell repair, although they will also provide energy for the body if eaten in excess of basic nutritional requirements.

These are just some of the reactions that are going on in your body, but there are many more situations in which chemical reactions occur. The chemist in industry plays a vital role ensuring that processes are economically viable, both in terms of production yields and time taken. Every reaction is different; it is necessary to be able to distinguish between them, to understand the optimum conditions for effective outcomes, and to be able to measure how quickly each one proceeds. The following programme has been written to develop your background knowledge and understanding of these basic ideas.

Energy Changes in Chemical Reactions

When a piece of magnesium metal is burned in air it produces a lot of heat and a brilliant white light; in older cameras this was seen when the 'flashgun' was used. The same reaction is now used to provide bright light in emergency 'flares' and fireworks.

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

When a mixture of hydrogen and oxygen gases is ignited the noise heard varies from a 'squeaky pop' to a deafening bang depending on the quantities present.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

If a small piece of sodium is placed in a bowl of water, it whizzes around the surface and melts into a silvery globule, it may even explode and burn with a yellow flame.

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

In these examples, some of the chemical energy that was stored within the bond structure of the reactants is given out during the reaction as heat or light, and even as movement or sound. All reactions involve energy changes and often energy is released, as heat or light for example, and more rarely some reactions absorb energy during their course. This unit considers various chemical reactions in which there is a noticeable or measurable amount of energy released or absorbed; it also provides a theoretical basis on which to build a picture as to why reactions have such different outcomes.

Useful website of video of combustion reactions including Mg and O_2.

http://www.rsc.org/education/chemistryteachers/Index_Results.asp?Page=1&ID=11265&Search

The heating of copper II sulphate and shifting the reaction between The hydrated and dehydrated forms can be used to illustrate endothermic and exothermic processes.

Activity 1: Heating copper(II) sulfate

Topic
Reversible reactions.

Timing
30 mins.

Description
Students remove the water of crystallisation from hydrated copper(II) sulfate by heating. Condensing in a test-tube collects the water. The white anhydrous copper(II) sulphate can then be rehydrated, the blue colour returns.

Apparatus and equipment (per group)
Two test-tubes
Delivery tube (right angled)
250 cm^3 beaker for cold water bath
Bunsen burner
Clamp and stand.

Chemicals (per group)
Hydrated copper(II) sulfate (powdered)(**Harmful**).

Teaching tips
Ensure that the reaction test-tube is clamped at the end nearest the bung. Warn about and watch for ‘suck back’. Demonstrate how to lift the entire clamp stand and apparatus.

Background theory
Some reactions are easily reversible and others are not. Concept of water of crystallisation.

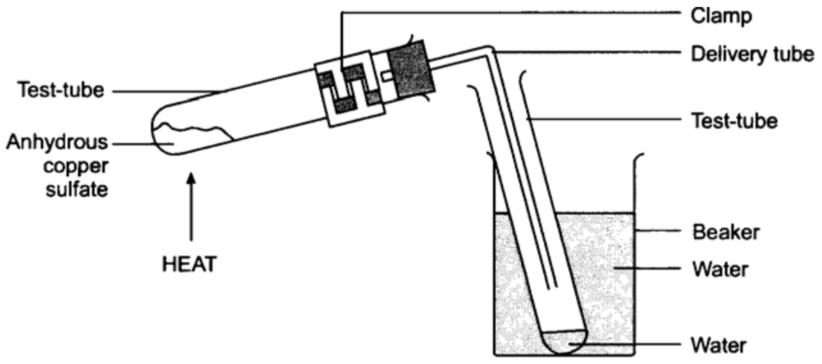
Health & Safety
Wear eye protection.

Answers
1. To cool and condense the steam.
2. *(a)* with water
   (*b*) without water
   (*c*) substance produced by a reaction
   (*d*) turned from vapour to liquid
   (*e*) a product is produced by the formation of new bonds.
3. The reaction can easily be reversed to produce the original reactants – *ie* hydrated copper(II) sulfate and heat.
4. The same amount of heat is required to form anhydrous copper(II) sulfate from the hydrated form, as that produced by the rehydration.

Heating copper(II) sulfate

Introduction

In this experiment the water of crystallisation is removed from hydrated blue copper(II)sulfate. After cooling the anhydrous copper(II) sulfate formed is then rehydrated with the same water.



What to record

Record any observations when the water was poured back onto the white copper(II)sulfate.

What to do

1. Set up the apparatus as shown.
2. Heat the blue copper(II) sulfate until it has turned white.
3. Act quickly to prevent suck back. Lift the clamp stand so that the delivery tube does not reach into the water in the test-tube.
4. Allow the anhydrous copper(II) sulfate to cool.
5. Hold the tube containing anhydrous copper(II) sulfate in one hand and pour the condensed water onto the powder.

Health & Safety

Wear eye protection.

Questions

1. Why is one test-tube placed in a beaker of cold water?
2. What do the following words mean (*a*) hydrated, (*b*) anhydrous, (*c*) product, (*d*) condensed and (*e*) reaction?
3. The reaction
Hydrated copper(II) sulfate + heat anhydrous copper(II) sulfate + water is called a reversible reaction. Why?
4. Anhydrous copper(II) sulfate could be used as a fuel for heating ('just add water to get the heat'). Explain why it would not be a very economical fuel.

Have you ever thought about:

Why magnesium burns so fiercely when lit? (After all, some alloys used in building aircraft contain magnesium and they do not burst into flames, unless they are heated very *strongly.)*
The significance of the flame or spark in the reaction between hydrogen and oxygen?
Why sodium reacts so spectacularly with water whereas iron rusts quite slowly?

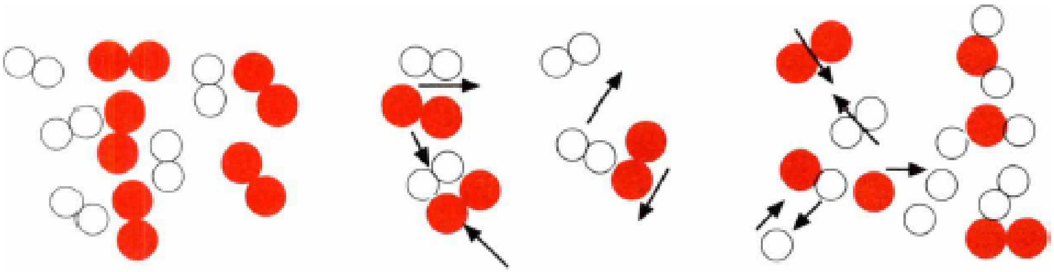
Collision Theory

All matter is made up of particles, and in the study of reaction kinetics we are interested in the behaviour of atoms, molecules and ions and how they react together. It seems very obvious to say this, but two chemicals can only react if their particles come into contact with each other.

The particles in gases and liquids are in constant motion and consequently they will collide with each other. Millions of such collisions occur every second, but not everyone produces a new substance; if they did reactions would be over in less than a millionth of a second. We have much everyday evidence to show that this is not so; for instance, if iron rusted on its first contact with the air then bridges and towers could never be built, so evidently not all collisions are effective.

Imagine a playground full of young school children. They are more likely to bump into each other when they are running around than when they are walking or sitting. Sometimes they will run past each other, sometimes they may just brush sleeves and stumble slightly, but occasionally two children will collide head-on! The difference in outcomes is both obvious and painful: the more children in the playground, the greater the likelihood of a collision. Likewise, in a chemical reaction the type of collision that occurs between particles, and how 'hard' or energetic it is, will help decide whether or not a reaction will happen. Also, just as in the overcrowded playground, the, more particles there are in a given volume - the greater their concentration - the greater the chance of collision.

We can use a simple reaction to illustrate this. Hydrogen and oxygen are both diatomic gases. Under the right conditions they can be made to react to form water. The diagram illustrates the way in which different molecules might collide.



The molecules move about randomly and at different speeds. Some may collide.

Some molecules only catch 'glancing blows' – low energy collisions with too little energy to break bonds. So no reaction!

'Head-on' collisions, real 'knock-outs' – break bonds in the hydrogen and oxygen molecules, allowing a new product to be formed.

Different types of collision in a mixture of gases

When two particles collide, 'head-on', with sufficient speed, the collision is a 'high energy' one and the reaction occurs quickly. Imagine two sports cars in a head-on collision! For instance, when the temperature is increased for a reacting system, the particles, on average, move faster - they have a greater kinetic energy (movement energy) and collisions are 'harder'. So, increasing the temperature increases the rate at which reactions occur.

SUMMARY

In order for a reaction to occur particles must collide with each other-the more particles that are in a certain amount of space, that is a fixed volume of gas or solution, and the faster they are moving, then the greater the likelihood of collisions and so a reaction. Increases in both temperature and concentration increase the rate of reaction: temperature has the greater effect.

Making and Breaking Bonds

All reactions involve the breaking of bonds, between the atoms or ions in the reactants, and the formation of new bonds to form different substances, *i. e.* the products. Atoms, ions and molecules involved in chemical reactions have two types of energy associated them:

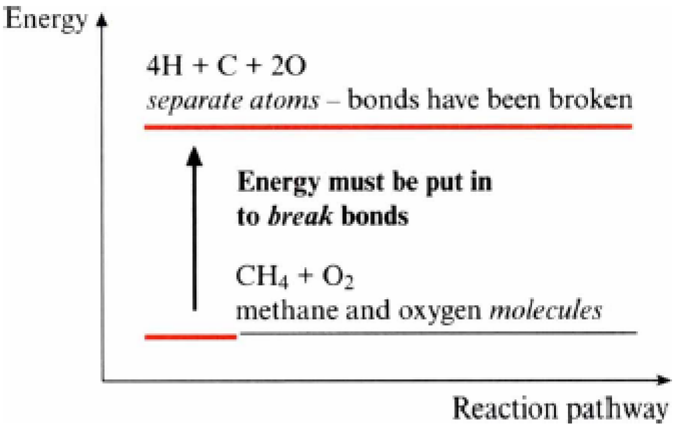
- **kinetic energy** - because they are moving;
- **bond energy** - because there are forces between the atoms or ions; these forces are the bonds that hold the atoms or ions together.

Energy is needed to break particles apart - to break bonds or overcome attractions. This energy is usually supplied when particles collide. However, if this is insufficient then more energy must be transferred somehow, and this is usually in the form of *heat.*

Strong bonds will need more energy to break them than weak bonds.

Coal and wood are both well known as domestic fuels. They do not burn without help, however - they both must be heated to fairly high temperatures before they will 'catch fire'. Only then can the chemicals in the fuels react with the air. Once started, the reaction continues, giving out much more heat, which in turn makes more fuel hot enough to ignite. The original input of heat has to be supplied in order to break apart the bonds between atoms of oxygen molecules, O_2, in the air, and the carbon and hydrogen atoms, *etc.,* that make up the chemicals in wood or coal. This initial supply of heat energy has been taken in, *i.e.* absorbed by the reactants. However, we are well aware that both coal and wood produce large amounts of heat when burned. This *excess heat* is given out as new bonds are formed in the making of carbon dioxide and water *etc.* as the products of combustion.

The energy diagram represents what is happening during the reaction as bonds are broken when methane gas *(e.g.* as ordinary North Sea gas used for cooking) is burned in air. It can be seen that the energy content of the system increases - bond breaking needs energy.



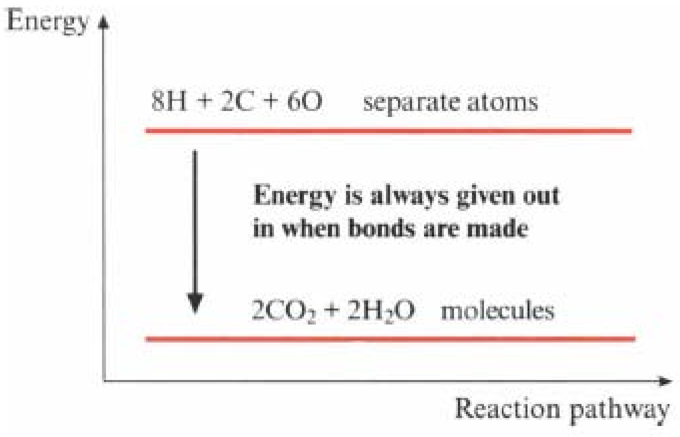
Energy diagram for the breaking of bonds in the oxidation of methane

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

The next stage of the reaction involves the making of new bonds as carbon dioxide and water are formed. It has already been said that, 'the stronger the bond the greater the energy needed to break it'; conversely, when bonds are formed, energy is *released:*

The stronger the bond formed the more energy is released - making bonds gives out energy.

The energy diagram for the formation of new bonds is shown below. It can be seen that the energy content of the system falls as the bonds are made.



Energy diagram for the formation of new bonds in the oxidation of methane

Exothermic and Endothermic Reactions

Chemical changes are always accompanied by *changes* in the energy content of the materials that are reacting, and the *change* is usually observed in the form of heat. Indeed, in many cases the change in temperature when substances react is the only evidence that a chemical reaction has taken place. Changes in the heat content of a system - called the heat of reaction - can be measured by the change in the temperature of the system during the course of the reaction. The symbol for 'heat of reaction' is ΔH (where Δ is the Greek letter *delta;* used in mathematics to mean 'difference').

EXOTHERMIC REACTIONS

When magnesium is burned in oxygen it burns with a white hot flame.

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s) \qquad \Delta H \text{ is negative}$$

The stored chemical energy in the magnesium and oxygen is partly transferred to stored chemical energy in the bonds of magnesium oxide while the rest is emitted as heat and light. Thus, the magnesium oxide has a lower overall energy content than the elements from which it is made.

When magnesium ribbon is added to dilute hydrochloric acid, you can feel the test tube getting hot! Heat is given out.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g) \qquad \Delta H \text{ is negative}$$

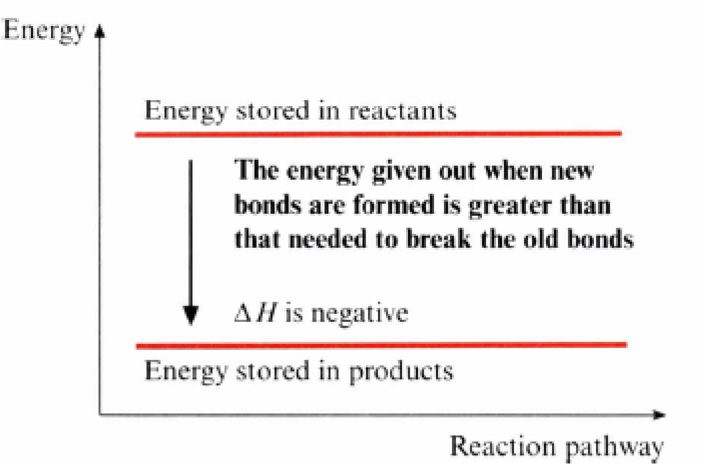
When water is added to white anhydrous copper sulfate powder, to form the hydrated blue solid, the reaction gives out heat and there is a rise in temperature until enough water has been added and the reaction stops.

$$\underset{\text{white}}{CuSO_4(s)} + 5H_2O(l) \rightarrow \underset{\text{blue}}{CuSO_4{\cdot}5H_2O(s)} \qquad \Delta H \text{ is negative}$$

All the reactions above give out energy to their surroundings, in particular, heat energy.

- Reactions that give out heat energy are called exothermic reactions.
- By *convention ΔH* for an exothermic reaction is given a negative sign, showing that heat is given out - *i.e.* heat energy has been lost from the reacting system to the surroundings - the system ends up with a *lower* energy content.

The general energy diagram for an exothermic reaction is shown below.



Energy diagram for an exothermic reaction

The reactions above all demonstrate *energy changes.*

The Law of Conservation of Energy states that 'energy can neither be created nor destroyed - it only *changes* from one form to another'.

This process can also be shown as a word equation for an exothermic reaction.

Energy stored in the bonds of of the reactants = Energy stored in the bonds of products + HEAT energy given OUT

A huge number and variety of chemical reactions are exothermic and have energy diagrams as illustrated below in some common exothermic reactions:

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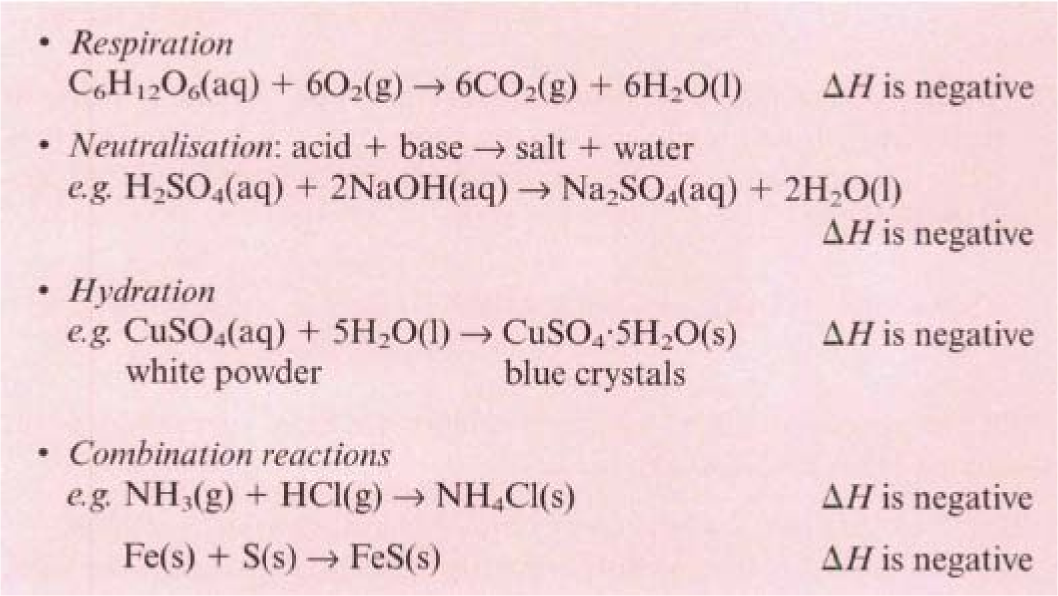


- *Respiration*
  $C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$ ΔH is negative
- *Neutralisation*: acid + base → salt + water
  e.g. $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
  ΔH is negative
- *Hydration*
  e.g. $CuSO_4(aq) + 5H_2O(l) \rightarrow CuSO_4{\cdot}5H_2O(s)$ ΔH is negative
  white powder blue crystals
- *Combination reactions*
  e.g. $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$ ΔH is negative
  $Fe(s) + S(s) \rightarrow FeS(s)$ ΔH is negative

ENDOTHERMIC CHANGES

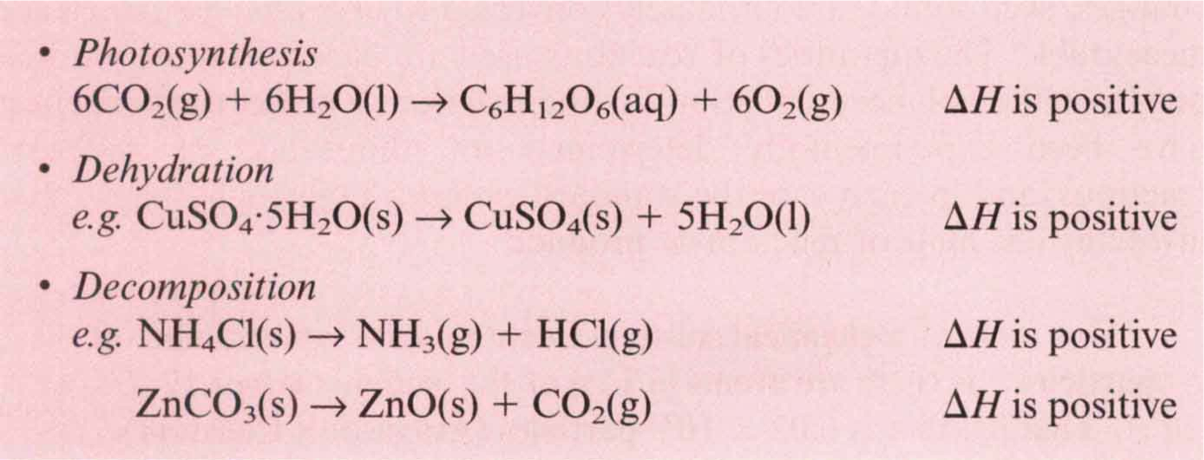
Look at what happens in the following processes:

- The decomposition of calcium carbonate to calcium oxide and carbon
- Some physical processes such as evaporation, dissolving and melting needs a lot of heat energy- they are endothermic processes.
- When ammonium nitrate is dissolved in water you can feel the test tube get colder as the dissolving process takes in heat from your hand and the surroundings.

So where does this absorbed heat energy go?

When salt is dissolved in water heat is absorbed from the surroundings. This heat energy is needed to overcome the forces of attraction between the sodium and chloride ions in the crystal lattice, so that they dissolve and are then free to move about in the solution.

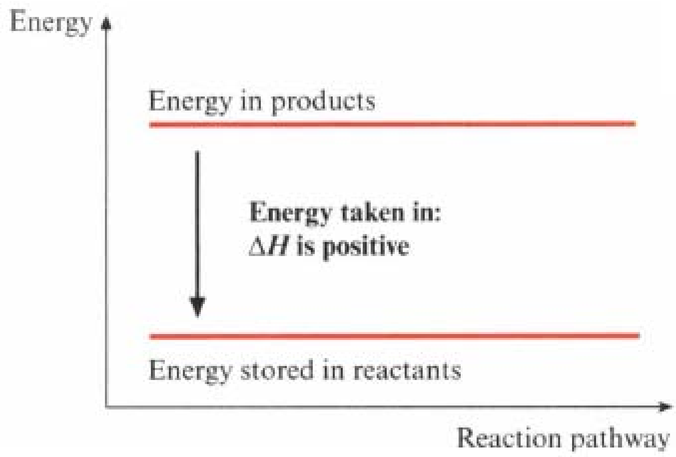
In each of the examples above, the energy content of the products is greater than that of the reactants. It is worth noting, however, that genuine, spontaneous, endothermic, chemical reactions are relatively rare, and some examples are given below as common endothermic reactions:



- *Photosynthesis*
  $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$ ΔH is positive
- *Dehydration*
  e.g. $CuSO_4{\cdot}5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(l)$ ΔH is positive
- *Decomposition*
  e.g. $NH_4Cl(s) \rightarrow NH_3(g) + HCl(g)$ ΔH is positive
  $ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$ ΔH is positive

When ENDOTHERMIC changes and reactions happen, heat is taken in from the surroundings.

By convention *ΔH* is given a positive sign for endothermic processes and reactions - *i.e.* heat energy is gained by the system from its surroundings and the system ends up with a higher energy content.



Energy diagram for an endothermic reaction

For an endothermic process this can be summarised by the word equation:

Energy stored in the bonds of of the products = Energy stored in the bonds of the reactants + HEAT energy taken IN

The following demonstrations, 'Endothermic solid-solid reactions' and 'Spontaneous exothermic reaction' can be effectively used to illustrate endothermic and exothermic reactions. The exothermic demonstration is also interesting as it is spontaneous requiring no additional heat input. This shows that the activation energy is ether at or below room temperature. It is useful if discussing this to compare the demonstration to that of burning magnesium. In that case heat has to be put into the system in order to supply the activation before a large amount of energy is released from the reaction.

Activity 2: Endothermic solid-solid reactions

Teacher notes

Solid hydrated barium hydroxide (Corrosive) is mixed with solid ammonium chloride (Harmful) in a beaker. An endothermic reaction takes place to produce a liquid, with the evolution of ammonia. The temperature drops dramatically to about -20 °C.

Lesson organisation

Although the experiment can be safely carried out as a class experiment (with GCSE or A-level candidates in mind), it lasts only about 10 minutes and is probably not worth the extra time spent by students setting up and clearing away. Therefore it is recommended as being more suitable as a teacher demonstration. Students could be allowed to feel the outside of the very cold container.

Procedure

HEALTH & SAFETY: Wear goggles

Before the demonstration

Weigh out separately the barium hydroxide and the ammonium chloride. Avoid lumps as far as possible.

The demonstration

Work in a fume cupboard unless the room is well ventilated (both products: ammonia gas and barium chloride are toxic).

a) Stand the beaker on a watch-glass containing a few drops of water, so that the base of the beaker is touching the water.
b) Note the room temperature.
c) Mix the two solids in the beaker and stir with the thermometer. The mixture becomes slushy as a liquid is formed, together with a white suspension.
d) The presence of ammonia can be detected by smell, and confirmed by blowing fumes from the hydrochloric acid (corrosive) bottle across the beaker's mouth and by using moist indicator paper.
e) Observe the drop in temperature, which is confirmed by the fact that the beaker freezes to the watch-glass.

Teaching notes

It helps to use a large thermometer display. The cold beaker can be passed around the class once the evolution of ammonia has stopped.

It is not possible to determine easily the exact barium compound or compounds produced in this reaction but the equation is usually represented as:

$Ba(OH)_2.8H_2O(s) + 2NH_4Cl(s) \rightarrow 2NH_3(g) + 10H_2O(l) + BaCl_2(s)$
or
$Ba(OH)_2.8H_2O(s) + 2NH_4Cl(s) \rightarrow 2NH_3(g) + 8H_2O(l) + BaCl_2.2H_2O(s)$

A-level students could be asked to calculate the value of the enthalpy and entropy changes for the reaction, using standard enthalpy changes of formation and standard entropy values obtained from a data book or from the table below.

Compound	$\Delta H_f°$/kJ mol^{-1}	S°/J mol^{-1} K^{-1}
$Ba(OH)_2.8H_2O(s)$	–3345	427
$NH_4Cl(s)$	–314	95
$NH_3(g)$	–46	192
$H_2O(l)$	–286	70
$BaCl_2(s)$	–859	124
$BaCl_2.2H_2O(s)$	–1460	203

An enthalpy change of +164 kJ mol^{-1} is obtained if the product is assumed to be $BaCl_2(s)$, and +135 kJ mol^{-1} if it is assumed to be $BaCl_2.2H_2O(s)$. Students should be able to predict qualitatively that the entropy change for the system has a positive value because a gas and a liquid are formed from two solids. From the values above they could also be asked to calculate the actual entropy change for the system and the surroundings, and hence ΔS_{total} or ΔG for the reaction and confirm that the process is spontaneous. A value of ΔS_{system} of +591 J mol^{-1} K^{-1} is obtained if the product is assumed to be $BaCl_2(s)$ and +530 J mol^{-1} K^{-1} if it is assumed to be $BaCl_2.2H_2O(s)$.

Web Links

This weblink shows a related endothermic reaction involving barium hydroxide :

http://jchemed.chem.wisc.edu/JCESoft/CCA/CCA3/MAIN/ENDO2/PAGE1.HTM

Endothermic solid-solid reactions

Technician notes

Apparatus and chemicals

One demonstration will require:

Eye protection: goggles
Beaker (100 cm^3)
Watch-glass
Thermometer, reading to -30 °C (see note 1)
Top-pan balance
Fume cupboard (optional) - Products ammonia gas See CLEAPSS Hazcard 5 toxic and barium chloride See CLEAPSS Hazcard 10A toxic

Barium hydroxide-8-water (Corrosive), 32 g
Ammonium chloride (Harmful), 10 g

Concentrated hydrochloric acid (Corrosive) (see note 2)
Universal indicator (or litmus) paper, 1 strip

Technical notes

Barium hydroxide-8-water (**Corrosive**) Refer to CLEAPSS Hazcard 10B
Ammonium chloride (**Harmful**) Refer to CLEAPSS Hazcard 9A
Concentrated hydrochloric acid (**Corrosive**) Refer to CLEAPSS Hazcard 47A

1 Consider using a thermocouple-type of thermometer which can be connected to a large display or computer monitor.

2 Small stock bottle, to provide fumes for ammonia test.

Activity 3: Spontaneous exothermic reaction

In this demonstration experiment, a mixture of glycerol (propane-1,2,3-triol) and potassium manganate(VII) crystals bursts into flame, giving off clouds of steam, after a short time lag.

Lesson organisation

This reaction can be used as a fun demonstration to show a spontaneous reaction, or as an example of the redox reaction between a fuel and a powerful oxidising agent. The time lag illustrates the speeding up of an initially slow exothermic reaction as the energy given out raises the temperature of the mixture.

[WARNING: The reaction can sometimes be delayed. I nothing happens, rinse the mixture down the sink with plenty of water]

Apparatus and chemicals

Eye protection
Safety screens
The teacher requires:
Clean metal lid from a tin can or jar (see note 1)
Heat resistant mat
The quantities of chemicals given are for one demonstration.
Access to:
Potassium manganate(Vll) (Oxidising agent, Harmful, Dangerous to the environment),
2-3 g in the form of fine crystals (see note 2)
Glycerol (propane-1,2,3-triol) (Low hazard), about 1 cm3 in a test-tube (see note 3)

Technical notes

Potassium manganate(VII) (Oxidising agent, Harmful, Dangerous for the environment)
Refer to CLEAPSS Hazcard 81
Glycerol (propane-1,2,3-triol) (Low hazard) Refer to CLEAPSS Hazcard 37

1. If the lid from a jar has a plastic lining, this should be scraped off. Alternatively, use a small foil cake case which has been cleaned and dried.
2. Fine crystals of potassium manganate(VII) work much better than larger ones. Use a pestle to grind large crystals in a clean mortar, if necessary.
3. Old samples of glycerol are sometimes ineffective, possibly because of absorbed water from the air.

The following class practical can be used to explore whether a reaction is exothermic or endothermic by observation.

Activity 4: Exothermic or endothermic?

Teacher notes

This is a useful class practical to introduce energy changes in chemical reactions. The students measure the temperature changes in four reactions, and classify the reactions as exothermic or endothermic. The experiments can also be used to revise different types of chemical reaction and, with some classes, chemical formulae and equations.

Lesson organisation

There are five solutions and three solids involved. Careful consideration will need to be given as to the most appropriate way to dispense these to the class. Special care should be taken with the magnesium ribbon and magnesium powder and, with some classes, teachers may prefer to dispense these materials directly.

The length of time required for carrying out the actual reactions is around 30 minutes, but this will depend on the nature of the class and how the practical is organised.

Procedure

HEALTH & SAFETY: Wear eye protection throughout.

Reaction of sodium hydroxide solution and dilute hydrochloric acid

a Stand the polystyrene cup in the beaker.

b Use the measuring cylinder to measure out 10 cm^3 of sodium hydroxide solution and pour it into the polystyrene cup.

c Measure the initial temperature of the sodium hydroxide solution and record it in a suitable table.

d Measure out 10 cm^3 of hydrochloric acid and carefully add this to the sodium hydroxide solution in the polystyrene cup. Stir with the thermometer and record the maximum or minimum temperature reached.

e Work out the temperature change and decide if the reaction is exothermic or endothermic.

f Discard the mixture (in the sink with plenty of water). Rinse out and dry the polystyrene cup.

Reaction of sodium hydrogencarbonate solution and citric acid

a Repeat steps **a** – **c** of the previous experiment, using sodium hydrogencarbonate solution in place of sodium hydroxide solution.

b Add 4 small (not heaped) spatula measures of citric acid. Stir with the thermometer and record the maximum or minimum temperature reached.

c Work out the temperature change and decide if the reaction is exothermic or endothermic.

d Discard the mixture (in the sink with plenty of water). Rinse out and dry the polystyrene cup.

Reaction of copper(II) sulfate solution and magnesium powder

a Repeat steps **a** – **c** of the first experiment, using copper(II) sulfate solution in place of sodium hydroxide solution.

b Add 1 small (not heaped) spatula measure of magnesium powder. Stir with the thermometer and record the maximum or minimum temperature reached.

c Work out the temperature change and decide if the reaction is exothermic or endothermic.

d Discard the mixture (in the sink with plenty of water). Rinse out and dry the polystyrene cup.

Reaction of sulfuric acid and magnesium ribbon

a Repeat steps **a** – **c** of the first experiment, using sulfuric acid in place of sodium hydroxide solution.

b Add one 3 cm piece of magnesium ribbon. Stir with the thermometer and record the maximum or minimum temperature reached.

c Work out the temperature change and decide if the reaction is exothermic or endothermic.

d Once all the magnesium ribbon has reacted, discard the mixture (in the sink with plenty of water). Rinse out and dry the polystyrene cup.

Teaching notes

The reactions and types of reaction involved are:

Sodium hydroxide + hydrochloric acid → sodium chloride + water (Neutralisation)

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

Copper(II) sulfate + magnesium → magnesium sulfate + copper (Displacement, Redox)

$CuSO_4(aq) + Mg(s) \rightarrow MgSO_4(aq) + Cu(s)$

Sulfuric acid + magnesium → magnesium sulfate + hydrogen (Displacement, Redox)

$H_2SO_4(aq) + Mg(s) \rightarrow MgSO_4(aq) + H_2(g)$

At this level the neutralisation reaction between sodium hydrogen carbonate and citric acid may be a bit complicated – it may be better to just use the word equation. More able students could use $H^+(aq)$ to represent the acid.

Sodium hydrogencarbonate + citric acid → sodium citrate + water + carbon dioxide

$NaHCO_3(aq) + H^+(aq) \rightarrow Na^+(aq) + H_2O(l) + CO_2(g)$

Exothermic or endothermic?

Technician notes

Apparatus and chemicals

Eye protection

Each group of students will need:

Polystyrene cup (expanded polystyrene)
Beaker (250 cm^3) in which to stand the polystyrene cup for support (see note 1)
Thermometer (–10°C to 110°C)
Measuring cylinder (10 cm^3), 2
Spatula
Absorbent paper

Access to the following solutions:
(all at approx 0.4 mol dm^{-3} concentration); (see note 2)

Copper(II) sulfate (**Low hazard**)
Hydrochloric acid (**Low hazard**)
Sodium hydrogencarbonate (**Low hazard**)
Sodium hydroxide (**Irritant**)
Sulfuric acid (**Low hazard**)

Access to the following solids (see note 3):

Magnesium ribbon (**Highly flammable**), cut into 3 cm lengths.
Magnesium powder (**Highly flammable**).
Citric acid (**Irritant**).

Technical notes

Copper(II) sulfate (solution - **Low hazard,** solid - **Harmful**) Refer to CLEAPSS Hazcard 27C and CLEAPSS Recipe Card 19
Hydrochloric acid (**Low hazard**) Refer to CLEAPSS Hazcard 47A and CLEAPSS Recipe Card 31
Sodium hydrogencarbonate (**Low hazard**) Refer to CLEAPSS Hazcard 95C and CLEAPSS Recipe Card 64

Sodium hydroxide (**Irritant**) Refer to CLEAPSS Hazcard 91 and CLEAPSS Recipe Card 65
Sulfuric acid (**Low hazard**) Refer to CLEAPSS Hazcard 98A and CLEAPSS Recipe Card 69
Magnesium ribbon (**Highly flammable**) Refer to CLEAPSS Hazcard 59A
Magnesium powder (**Highly flammable**) Refer to CLEAPSS Hazcard 59A
Citric acid (**Irritant**) Refer to CLEAPSS Hazcard 36C

1 Typical expanded polystyrene cups fit snugly into 250 cm^3 squat form beakers. This provides a more stable reaction vessel and also prevents spillage if the polystyrene cup splits.

2 At the suggested concentrations, the solutions (except for sodium hydroxide) represent minimal hazards, although it is probably advisable to label them as **Harmful**. If the concentrations are increased then the solutions must be labelled with the correct hazard warning. The solutions could be provided in small (100 cm^3) labelled conical flasks or beakers.

3 Small amounts of the solids can be provided in plastic weighing boats or similar. The teacher may prefer to keep the magnesium ribbon and powder under their immediate control and to dispense on an individual basis.

Heat of Reaction

As stated above, the amount of energy given out or absorbed during a chemical reaction is called the heat of reaction: *ΔH.*

$$\textbf{The heat of a reaction} = \begin{array}{l}\textbf{Energy content of the products} \\ \quad - \textbf{Energy content of the reactants}\end{array}$$

In order to compare the amounts of energy gained or lost during a chemical reaction, it is necessary to have an agreed standard. The energy changes are measured as changes in the heat content of the reacting system, and so it is important that no energy is 'lost' in any other form.

Experiments to measure heats of reaction are carried out in sealed containers surrounded by a known volume of water, so that all energy changes, even sound, are ultimately converted to heat changes which are measurable. The quantities of reactants used are based on those worked out from the balanced equation for the reaction. The heats of reaction have been experimentally determined for thousands of different reactions, and in each case the standard energy change is based on that involving one mole of reactant or product.

One mole of a chemical substance contains the same number of particles as there are atoms in 12 g of the isotope carbon-12 (^{12}C).

That number is 6.02 x 10^{23} particles (Avogadro's Constant).

For instance, the reaction between nitrogen and hydrogen to produce ammonia can be represented by the equation:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92 \text{ kJ mol}^{-1}$$

But what is the heat of reaction for? One mole of ammonia or hydrogen or nitrogen?

In order to be consistent we must state the heat of reaction with reference to the whole stoichiometric equation - so the standard heat of reaction for the production of one mole of ammonia is *ΔH* = -46 kJ mol-l, based on:

$$\tfrac{1}{2}N_2(g) + \tfrac{3}{2}H_2(g) \rightleftharpoons NH_3(g) \quad \Delta H = -46\ \text{kJ mol}^{-1}$$

where mol-1 means ‘per molar quantities indicated by the equation’. The stoichiometric equation has been ‘re-balanced’ to show the molar quantities of hydrogen and nitrogen needed to produce one mole of the product ammonia.

Taking the earlier example of burning methane gas in oxygen:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -890\ \text{kJ mol}^{-1}$$

When methane burns in air, one mole of methane combines with *two moles* of oxygen and 890 kilojoules (kJ) of heat is evolved.

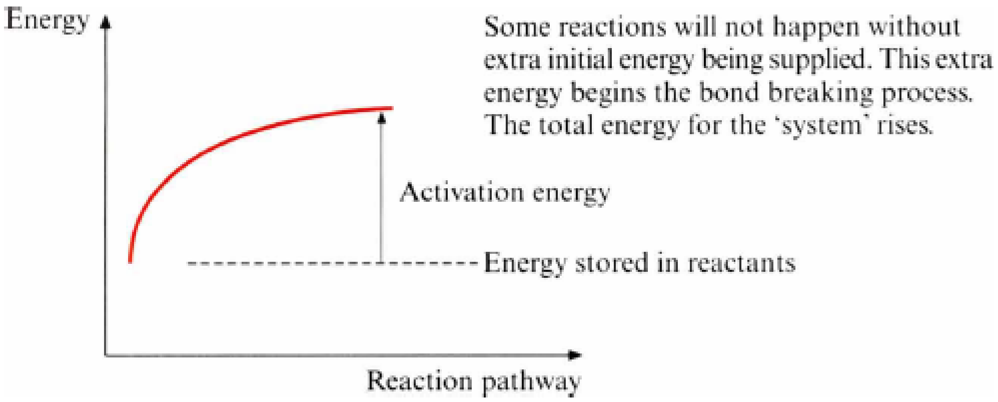
This is a *combustion reaction* and the heat of reaction, *ΔH,* is called the heat of combustion, ΔH_c, and is the amount of energy given out when one mole of methane burns in oxygen. In this case we would refer to the heat of combustion of methane, *i.e.* the heat change when one mole of methane is burned.

In practice, the data for heats of reaction are stated with reference to standard conditions of temperature and pressure, *i.e.* 298 K and 101.325 kPa (25 ^{o}C and 1 atm). If heats of reaction are measured under different conditions of temperature and pressure, it is likely that they will differ from this data.

Activation Energy

But why do some reactions occur ‘spontaneously’ whereas others need an energy boost?

We have seen that, for a chemical reaction to happen, molecules need not only to collide but must do so with sufficient energy. Breaking bonds needs energy; if molecules are moving too slowly and the collisions are weak, then there may not be sufficient energy to break the bonds and the reaction cannot proceed. The course of a chemical reaction might be compared to a pole-vault competition: only some competitors will have enough energy to get over the bar - others will need extra assistance or even the lowering of the bar, and some may never make it! The pole vaulter may improve performance by the use of technology to design a better pole or by drinking ‘high energy’ products. The equivalent may be achieved in a reaction by increasing the kinetic energy of the particles - making them move faster so that more of them have sufficient energy to get over the barrier. This can be done by heating the reactants.



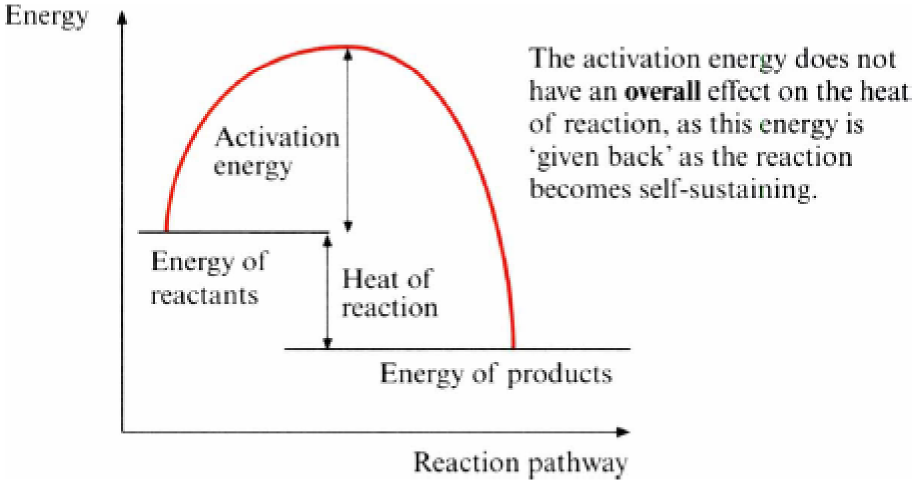
The energy barrier for a reaction - the activation energy

The minimum kinetic energy that particles must have before a reaction can take place is called the activation energy.

The activation energy, *E*,, is the minimum kinetic energy needed to enable bonds to stretch and break so that atoms, ions or molecules can rearrange as the reaction proceeds and so form products.

More simply, it is the energy needed to break the *old* bonds before the *new* bonds can form.

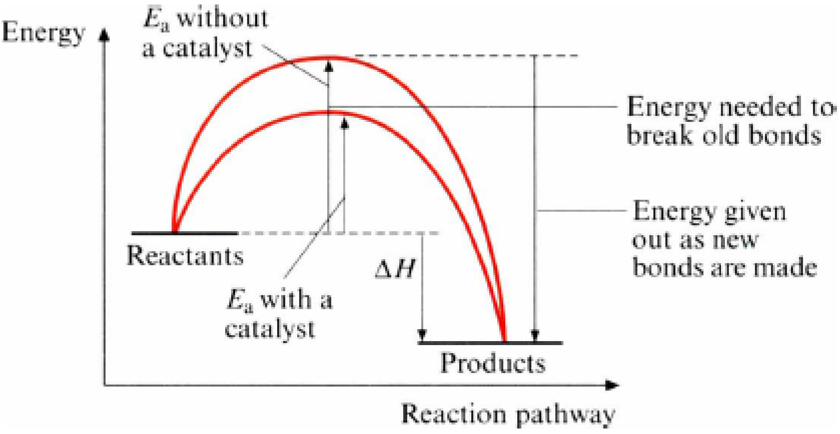
Reactions that occur readily at room temperature have a relatively low activation energy, whereas those that need large inputs of heat have high activation energies. Increasing the temperature of a reacting system raises the kinetic energy of the particles so there are more high impact collisions which then result in a reaction.



Energy diagram with summary of terms for a reaction

Some reactions can be 'sped up' by the use of catalysts. It is thought that the catalyst provides an alternative reaction pathway (mechanism) which has a lower activation energy.

This could also be shown as below:



Speeding up a reaction by use of a catalyst

It can be seen that when an exothermic reaction occurs, the new bonds are stronger than the old ones. When an endothermic reaction occurs, the new bonds are weaker than the old ones.

A class practical that illustrates the heat of energy is that of the combustion of alcohols. This allows for the comparison of different alcohols, different chain lengths, and the calculation of the energy emitted. It can be used along the same lines as a piece of coursework giving the students practice in the collection, analysis and evaluation of data as well as evaluating practical design and limitations.

Activity 5: Measuring heat energy of fuels

The combustion of alcohols is exothermic and in this experiment the energy released from burning a known mass of alcohol is used to heat a known amount of water. A comparison of various alcohols as fuels can be made by calculating the quantity of energy transferred to the water.

Lesson organisation

Students could undertake this experiment in pairs or small groups, perhaps using a different alcohol each and then pooling their results. Alternatively they could obtain their own results for each alcohol if time allows.

Burners containing the alcohols (or dropper bottles) need to be prepared in advance (see Technicians' notes).

Apparatus

Eye protection

All students will need access to:

Balances (2 or 3 d.p.)

Each student or pair of students will need:

Eye protection for each student
Clamp, stand and boss
Metal can (such as a copper calorimeter) or conical flask (250 cm^3)
Measuring cylinder (100 cm^3)
Spirit burner with cap or small crucible with wide lid (see note 1 to 4 and alternative method in teaching notes)
Stirring thermometer (0-110 °C)
Glass rod
Heat resistant mat

Access to

Matches or Bunsen burner (see note 6)
Balance (see note 7)

All students will need access to:

Alcohols in labelled spirit burners (or in dropper bottles if using crucibles) (see note 5):
Methanol (**Highly flammable, Toxic**)
Ethanol (**Highly flammable**)
or Industrial Denatured Alcohol (IDA) (**Highly flammable, Harmful**)
Propan-1-ol (**Highly flammable, Irritant**)
Butan-1-ol (**Flammable, Harmful**)

Technical notes

Methanol (**Highly flammable, Toxic**) Refer to CLEAPSS Hazcard 40B
Ethanol (**Highly flammable**) and Industrial Denatured Alcohol (**Highly flammable, Harmful**) Refer to CLEAPSS Hazcard 40A
Propan-1-ol (**Highly flammable, Irritant**) Refer to CLEAPSS Hazcard 84A
Butan-1-ol (**Flammable, Harmful**) Refer to CLEAPSS Hazcard 84B

1 If possible, provide spirit burners which have removable glass caps or, if using crucibles, provide lids. This makes it easier and safer to extinguish the flames.

2 Make sure the wick fits tightly in the wick holder and that the wick holder fits tightly in the burner.

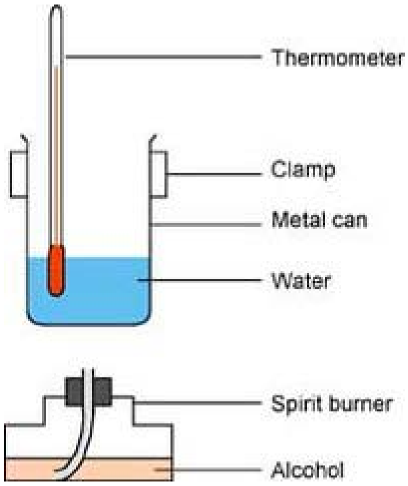
3 If burner is large capacity, reduce it by partly filling with epoxy resin or packing with cotton wool.

4 Fill and label spirit burners or dropper bottles with alcohols in advance of the lesson. Ensure any excess alcohol is wiped off the side of the burners or dropper bottles.

5 Keep bottles of alcohol well away from flames. Keep all stock bottles of alcohol in the prep room during the lesson.

6 Careful consideration must be given to how the students will ignite the alcohols. They could be given matches. Alternatively, a single Bunsen burner could be provided at the front of the lab (well away from the spirit burners or dropper bottles) and the students given access to wooden spills. The spirit burners *must* be kept upright when lighting. Do not tip onto the side.

7 The experiments should be carried out well away from the balances. Ensure, if possible, that students have access to balances at a number of different stations around the lab.



Procedure

HEALTH & SAFETY: Wear eye protection

a) Clamp the metal can (or flask) at a suitable height to allow room for the spirit burner to be placed below. Allow a gap of around 2-5 cm between the base of the flask/can and the top of the spirit burner. This gap may need to be adjusted depending on the height of the flame.
b) Using the measuring cylinder, fill the metal can with 100 cm^3 of water.
c) Using the thermometer, measure and record the initial temperature of the water.
d) Weigh the spirit burner (and cap) containing the alcohol and record the initial mass and name of the alcohol in a suitable table.
e) Place the spirit burner on the heat-resistant mat under the metal can, remove the cap, and light the wick.
f) Allow the alcohol to heat the water so the temperature rises *by* about 40°C. Use a glass rod or the thermometer to stir the water gently whilst the alcohol burns.
g) Replace the cap on the spirit burner to extinguish the flame.
h) Record the final temperature of the water using the thermometer. Work out the temperature change.
i) Reweigh the spirit burner and cap. Work out the mass of alcohol used.
j) Repeat the experiment for different alcohols using 100 cm^3 of fresh cold water each time.
k) If time allows repeat the experiment for each alcohol at least twice.

Teaching notes

Students should use their measurements to work out the temperature change and the mass of alcohol burned for each alcohol.

The energy transferred to the water from the burning alcohol can be calculated using the equation
$q = mc\Delta T$

where q = energy transferred (in J), m = mass of water (in g) c = specific heat capacity of water (in J/[g°C]) and ΔT = temp change (in °C or K).

Assume that 1 cm^3 of water has a mass of 1 g.
Assume that the specific heat capacity of water, c = 4.18 J/(g°C) or c = 4.18 J g^{-1} K^{-1} (i.e. 4.18 J are required for every 1 °C rise in temperature per g of water).

Students could be asked to work out the answer which fuel (alcohol) provides the most energy per gram of fuel burned?

The experimental results are often much less than values given in data books, or predicted using models and the energies needed to break the bonds. This is largely due to the fact that not all the energy produced is transferred to the water. Some is used to heat the air and surroundings, plus the flask. Also, combustion of the alcohols is likely to be incomplete. You may like to ask the students how they could tell this from their observations of the fuels burning.

Additional work: Students could investigate methods of reducing heat loss to the environment and thus discuss how the experiment could be improved.

For advanced courses, the experiment can be adapted to determine values for the enthalpy of combustion. The experimental set-up can be calibrated by using, say, propan-1-ol, and then the method used for other alcohols. This gives more accurate results.

Alternative method

If using small crucibles to burn the alcohol then students should:

a) Measure the mass of the crucible and, well away from naked flames, add about 0.5 g of the alcohol from the dropper bottle. Measure the mass again.
b) Place the crucible on the heat resistant mat under the conical flask and ignite the alcohol.
c) Stir the water in the flask with the glass rod or thermometer and record the highest temperature reached by the water.
d) Allow all the alcohol to burn away.

Energy in forms other than heat

So far, whilst discussing exothermic and endothermic reactions, we have been concentrating upon the measurement of heat to follow the reaction. However, there are other forms of energy that can be absorbed or emitted. Just think about an explosion, a truly exothermic reaction. It involves heat but also light and sound.

A good example of this is chemiluminescence. In this case the energy emitted by the reaction between the alkaline luminol and bleach solution is released a light in the blue region. This can be easily demonstrated using the following procedure:

Activity 6: Chemiluminescence – cold light

Teacher notes

Wear goggles.
A solution of sodium chlorate(I) oxidises an aqueous solution of luminol (3-aminophthalhydrazide). The reaction gives out a blue chemiluminescent glow without any increase in temperature of the mixture.

To view a video clip of this demonstration experiment, click here
(http://media.rsc.org/videoclips/demos/Chemiluminescence.mpg).

Lesson organisation

This demonstration experiment shows that a chemical reaction can give out energy as light instead of heating up its surroundings. The demonstration can also be used to stimulate interest in chemistry at an Open Day or other public event. It can also be undertaken as a group practical.

Apparatus and Chemicals

The teacher will require:

Eye protection
Conical flasks with stoppers (1 dm^3), 2
Beaker (2 dm^3)
Thermometer (0 – 100 °C)
Balance (1 d.p.)

Quantities of chemicals for one demonstration

Luminol (3-aminophthalhydrazide), 0.4 g (**Irritant**) (see note 4)
Sodium hydroxide pellets, 4.0 g (**Corrosive**)
Household bleach, 100 cm^3 (**Irritant**) (see notes 1, 2 and 3)
Fluorescein (or sodium fluorescein) powder (optional) (**low hazard)**

Technical notes

Luminol (**Irritant**) Refer to CLEAPSS Hazcard 4B
Fluorescein (**Low hazard**) Refer to CLEAPSS Hazcard 32
Sodium hydroxide pellets (**Corrosive**) Refer to CLEAPSS Hazcard 91
Sodium chlorate(I) solution (**Corrosive**) Refer to CLEAPSS Hazcard 89
Household bleach (**Irritant**) Refer to CLEAPSS Hazcard 89

1 Household bleach, which is typically a 5% 'available chlorine' solution, can be used. Make sure that the household bleach contains sodium chlorate(I) (sodium hypochlorite), NaOCl, and *not* hydrogen peroxide, as the bleaching agent. Many household bleaches nowadays also contain thickeners and/or detergents. Use ‘economy’ bleaches without any additives.

2 The sodium chlorate(I) solution sold by chemical suppliers contains up to 14% available chlorine. It has a limited shelf life. Adjust the volumes of bleach and water to make up the diluted bleach solution for the demonstration. See **Procedure** step **a**.

3 Tap water can be used for making up the solutions. The solutions are stable for over 12 hours and so can be made up well in advance.

4 The luminol does not always appear to dissolve completely, leaving a fine, greenish suspension.

Procedure

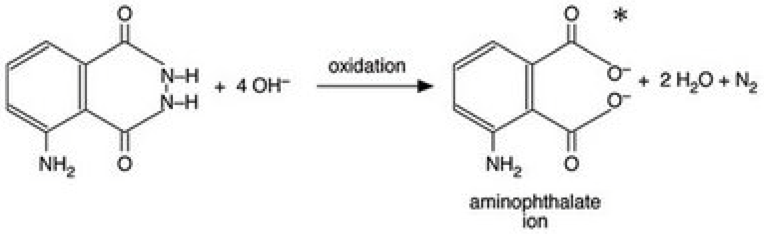
HEALTH & SAFETY: Wear eye protection.

a) Take the temperature of the solutions.
b) Lower the room lights and slowly pour the two solutions at the same rate into the beaker so that they mix. A pale blue glow will be seen which will last for a few seconds.
c) Take the temperature of the mixture. It will be the same as that of the starting solutions.

Teaching notes

For more dramatic effect when demonstrating, pour the solutions into a funnel attached to clear tubing bent into a variety of shapes, such as a spiral.

In the reaction, luminol is oxidised by the bleach to the aminophthalate ion, which is produced in an electronic excited state. This gives out energy as light (fluorescence) when it decays to the ground state.



Adding a small amount of fluorescein to the luminol solution, just before the demonstration, will alter the glow to a yellow-green colour.

Chemiluminescent 'light sticks' will be familiar nowadays to many students. A different reaction is used, involving the oxidation of a di-ester by hydrogen peroxide in an organic solvent. This reaction is much slower, the glow continuing for some hours. One solution is contained inside a glass phial inside a plastic tube containing the other solution. Bending the plastic tube breaks the glass phial, allowing the two reactants to mix. Colour effects are obtained by adding dyes to which the excitation energy is transferred, the excited dye molecules then emit light of different wavelengths. Can you slow the reaction down by placing the light stick in a freezer?

Other recipes are given on CLEAPSS Recipe Book sheet 23.

Web Links

Chemiluminescence (http://media.rsc.org/videoclips/demos/Chemiluminescence.mpg)

With the appropriate students and a risk assessment it is possible to undertake this as a class practical but on a much reduced scale.

.Useful websites

Energy from fuels – a document looking at the process of gaining energy from fuels and the implication of their combustion.

http://media.rsc.org/Whats%20your%20reaction/WYR3n.pdf

Diagnostic Test

1 Which reactions, in solution, are quicker, ionic or covalent? **(1)**

2 In each of the following test tube reactions how could you tell if a chemical reaction had occurred?

a) Silver nitrate solution and hydrochloric acid are added together.
b) Magnesium metal is added to hydrochloric acid.
c) Water is added to quicklime (calcium oxide). **(3)**

3 The rate of a chemical reaction increases when the temperature of the system is raised. Why is this? **(2)**

4 What is meant by the 'activation energy' of a reaction? **(2)**

5 What is meant by the terms 'exothermic' and 'endothermic'? **(2)**

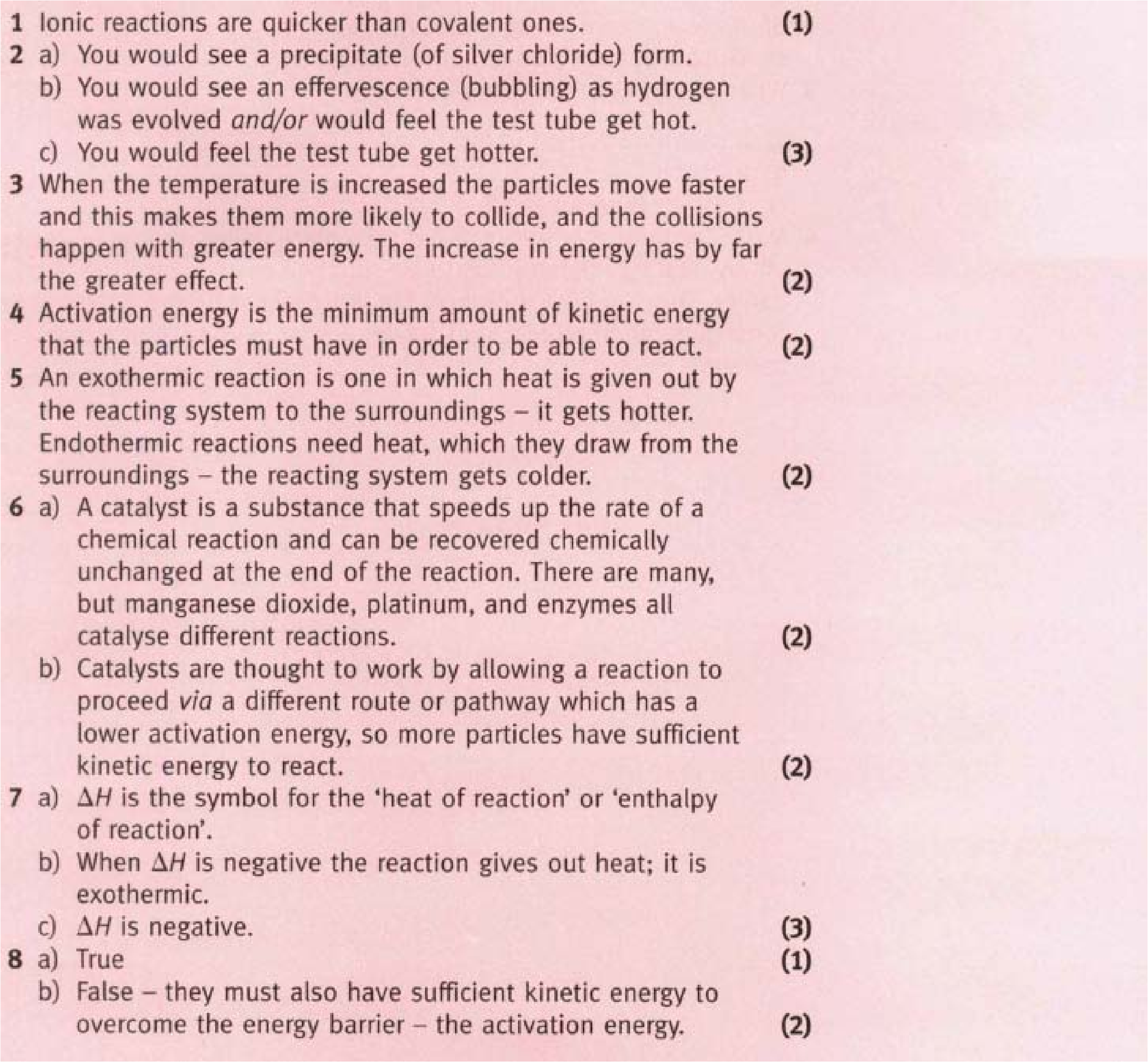
6 Some reactions are affected by catalysts.

a) Explain what a catalyst is and give an example. **(2)**
b) Describe briefly how catalysts are believed to work. **(2)**

7 a) What does the symbol ΔH represent?
b) What do you understand by 'ΔH is negative'.
c) Would ΔH be positive or negative for the reaction in Question 2(c)? **(3)**

8 a) More energy is released when strong bonds are formed than when weaker bonds are made. TRUE or FALSE? **(1)**
b) So long as particles collide, they will react together. TRUE or FALSE? Explain your answer. **(2)**

Answers



1 Ionic reactions are quicker than covalent ones. **(1)**

2 a) You would see a precipitate (of silver chloride) form.

b) You would see an effervescence (bubbling) as hydrogen was evolved *and/or* would feel the test tube get hot.

c) You would feel the test tube get hotter. **(3)**

3 When the temperature is increased the particles move faster and this makes them more likely to collide, and the collisions happen with greater energy. The increase in energy has by far the greater effect. **(2)**

4 Activation energy is the minimum amount of kinetic energy that the particles must have in order to be able to react. **(2)**

5 An exothermic reaction is one in which heat is given out by the reacting system to the surroundings – it gets hotter. Endothermic reactions need heat, which they draw from the surroundings – the reacting system gets colder. **(2)**

6 a) A catalyst is a substance that speeds up the rate of a chemical reaction and can be recovered chemically unchanged at the end of the reaction. There are many, but manganese dioxide, platinum, and enzymes all catalyse different reactions. **(2)**

b) Catalysts are thought to work by allowing a reaction to proceed *via* a different route or pathway which has a lower activation energy, so more particles have sufficient kinetic energy to react. **(2)**

7 a) ΔH is the symbol for the ‘heat of reaction’ or ‘enthalpy of reaction’.

b) When ΔH is negative the reaction gives out heat; it is exothermic.

c) ΔH is negative. **(3)**

8 a) True **(1)**

b) False – they must also have sufficient kinetic energy to overcome the energy barrier – the activation energy. **(2)**

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