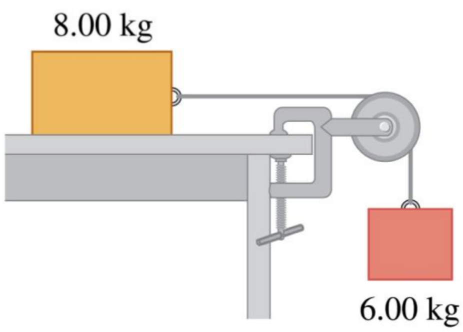
Exhothermic and Endothermic Reactions and Activation Energy

Let me start by explaining what ENERGY is.

Energy is the capacity to do work (or, to use the technical term, the **potential** to do work), and it can manifest itself in many different forms. An object that is above the floor has energy because if it is allowed to fall to the floor, then *it could do work*. For instance, if I were to tie[1] a piece of string to an object that is above the floor, and let the object go, then as it fell to the floor it would pull the object to which it was tied – and in moving the object **work would be done**. The suspended object thus has the potential to do work – and that particular energy is called "**gravitational potential energy**".



There are many stores of energy such as chemical, kinetic, gravitational potential, elastic potential and thermal stores, and energy can be transferred from one energy store to another by heat, by waves, by electric current or by a force moving an object. As it happens, (in Physics last Summer) I did a Note for you on Energy, and rather than repeat myself I have included that Note as an Appendix to this Note. Please make sure that you read and study that Note. You might also want to watch the following video at https://www.bbc.co.uk/bitesize/guides/z8hsrwx/revision/1
but the claim that energy can never be created is **incorrect** because there was nothing before the Universe came into being and there was something after the Universe came into being (presumably at 'Big Bang"), therefore at that moment **energy WAS created**. The Universe is made of matter, and matter cannot explain its own existence. The existence of 'matter' needs an explanation and some people call that creative force God/Allah/Jaweh/whatever. "Matter" as we know it must have come into existence at some point (presumably at 'Big Bang'). I myself do not have a big enough brain to understand a creative force, therefore all that I can do is to acknowledge my ignorance and then get on with things that I can understand. *Good scientists always make claims that* ***can be verified*** *rather than make unsubstantiated ones.*

OK, before getting on to AQA's text, let me also spend a bit of time explaining energy.

[1] By the way, for those of you whom I am teaching French, please notice that since this is a *hypothetical* example (I have not actually tied anything). I am merely hypothesising the tieing of something to something else, therefore I am using the Subjunctive Mood. In French the Present Subjunctive of the verb "attacher" would be "Que j'attache …." = "Were I to attach …". However, in modern French, one would use the Conditional Mood and say "Si je devais attacher ….." = if I were to attach ……

Energy is defined as the capacity to do work therefore it is a potential. It is just a capacity and as such energy cannot be seen, and when it is just a *potential* it is exceedingly difficult to measure energy. Luckily, when it has performed work, we can measure the work that was done, and **then we can say how much energy was used to do the work.**

It is difficult to measure an ***absolute*** amount of potential, but once work has been done then it is very easy to measure how much work **has actually been done**.

The next thing that we need to do is to familiarise ourselves with a new word: **Enthalpy**.

ENTHALPY

Enthalpy (for which the symbol is "**H**") is defined as

enthalpy = the internal energy of a system + (the product of its pressure times its volume)

$$\mathbf{H = E + PV}$$

I have already told you that it is very difficult to measure the *potential* for something (i.e. a capacity to do work), but that it is very easy to measure the amount of work that has *actually been done* (by measuring the change in Enthalpy that has taken place due to a chemical reaction), therefore in Chemistry we routinely measure the change in Enthalpy (ΔH) but not Enthalpy ("E") itself.

If the pressure and the volume of the system is kept constant i.e. not altered, then look what happens:

$$\text{change in enthalpy} = \Delta H = (E_2 + PV) - (E_1 + PV) =$$

$$\mathbf{\Delta H = E_2 - E_1}$$

and (because we kept the pressure and volume of the system constant) we eliminated the term "PV". We thus made the job of measuring Energy much easier.

In Chemistry (as opposed to say Physics), at GCSE Level, there are four standard ways of measuring work done

1) by using experimental data where you use the formula **"$\Delta H = m . c . \Delta T$"**
2) by using Enthalpy Profile Diagrams
3) by using Hess' Cycles, and
4) by using Average Bond Enthalpies/Bond Energy Terms

and I will explain each of the above terms when I come to the different Sections in Section 7 – but there are two more terms that I need to explain to you viz. "Exhothermic" and "Endothermic". (By the way. I still use the old-fashioned way of spelling "Exhothermic" i.e. with an 'h' in "Exho".)

EXHOTHERMIC vs. ENDOTHERMIC

The Term "**Exhothermic**" indicates that energy is being **given out** during the reaction concerned, while the Term "**Endothermic**" indicates that energy is being **used up/consumed** during the reaction concerned.

> The easiest way of remembering the difference between the two terms (Exhothermic and Endothermic) is to remember that *an "exit" is something where things leave*, therefore **in an "exhothermic" reaction, energy (usually in the form of heat) LEAVES the reaction**.

If there is a net release of energy during a reaction, then this fact is indicated by a negative sign; whereas, if there is a net consumption of energy during a reaction, then this is indicated by a positive sign viz.

- for **Exhothermic** reactions, ΔH has a " –ve " sign (energy is given out/released), and
- for **Endothermic** reactions, ΔH has a " +ve " sign (energy is consumed).

When the gas (Methane) in a cooker is turned on and is ignited, then the reaction that occurs is the combustion of the Methane (CH_4) with some of the Oxygen (O_2) that is in the air – and 890 kilojoules of energy are given off for every mole of CH_4 that is burnt.

$$\mathbf{CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O \qquad \Delta H = -890\ kJ\ mol^{-1}}$$

and, when the gas burns, it gives off heat (a*nd that is what cooks food*). The combustion of Methane is an **Exothermic** reaction.

Endothermic reactions are reactions that suck energy in. Probably the best-known (but least appreciated) Endothermic reaction of all is that of **Photosynthesis** where energy (from sunlight) is sucked in/is consumed/is used up when Carbon Dioxide and Water are converted into Glucose and Oxygen. This is an enormously important reaction in Biology and the reaction equation is as follows

$$\mathbf{CO_2 + H_2O \longrightarrow C_6H_{12}O_6 + O_2 \qquad \Delta H = +2{,}850\ kJ\ mol^{-1}}$$

Please note that the sign for an Endothermic reaction is now a positive ("+ve") one, whereas the sign for Exothermic reactions is negative ("–ve").

Actually, now that you are learning about the Chemistry of Exothermic and Endothermic reactions, you will be able to appreciate better something that you have learnt in Biology.

Photosynthesis and **Cellular Respiration** are essentially the same reaction except that they are the reverse of each other. In Photosynthesis energy from sunlight is used to convert Carbon Dioxide and Water into Oxygen and a Carbohydrate (usually Glucose), whereas in Cellular Respiration Oxygen and Glucose are being reacted together to produce energy (that is then stored in the bonds of ATP until the energy is need by the cell). If you have never appreciated this fact before, then I imagine that you will go round hugging yourself in delight at having made such an important (and beautiful) intellectual connection. Let me write in chemical symbols what I have just said in English words. (NB I have seen the ΔH value for this equation vary from 2,810-2,870 kJ mol^{-1}. It seems quite a large range, but perhaps ΔH values in Biology can very more than in Chemistry.)

Photosynthesis (an ENDOTHERMIC reaction)

			Sunlight					
Carbon Dioxide	**+**	**Water**	**—>**	**Oxygen**	**+**	**ENERGY (stored in glucose molecules)**	**; $\Delta H = +2{,}850$ kJ mol^{-1}**	
$6CO_2$	+	$6H_2O$	—>	$6O_2$		**ENERGY**	$C_6H_{12}O_6$	
6		6		6			1	Stoichiometric Ratios

Cellular Respiration (an EXHOTHERMIC reaction)

Glucose	**+**	**Oxygen**	**—>**	**ENERGY**	**+**	**Carbon Dioxide**	**+**	**Water ; $\Delta H = -2{,}870$ kJ mol^{-1}**	
$C_6H_{12}O_6$	+	$6O_2$	—>	(stored in ATP)		$6CO_2$	+	$6H_2O$	
1		6				6		6	Stoichiometric Ratios

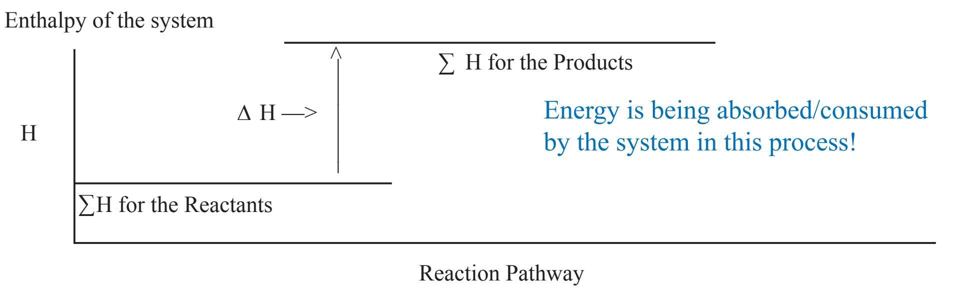
Cells need energy and the mitochondria in cells react Glucose and Oxygen together to produce the energy that is needed (and it is either used immediately or it is stored in the bonds of ATP), while plants use the energy in sunlight to convert Carbon Dioxide and Water into energy which is then stored in sugar molecules[3] and then distributed all over the plant via the Phloem System.

Can you see how Photosynthesis and Cellular Respiration are merely the reverse of each other, and that is why one is Endothermic and the other is Exothermic.

[3] NB Plants can manufacture Glucose (and other sugars), but animals **cannot** manufacture Glucose therefore they have to eat vegetable matter to obtain their glucose (and breathe in, from the air, the Oxygen that they need).

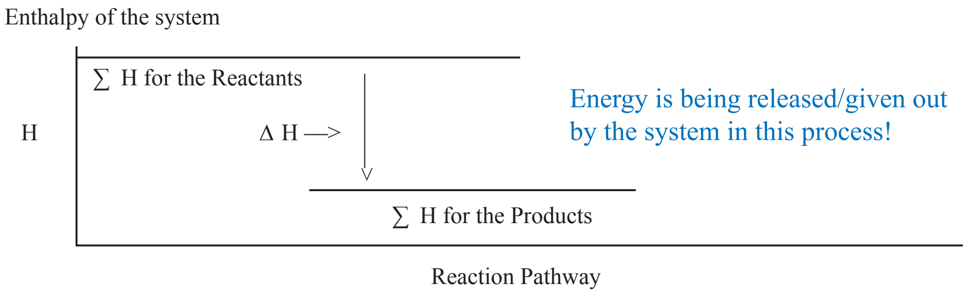
Energy is sucked in/consumed in an Endothermic Reaction therefore the products will contain more energy than the reactants, whereas in an Exhothermic reaction energy is released/given out therefore the products will contain less energy than the reactants. This can be seen clearly in the next two diagrammes. NB The symbol "Σ" means "the sum of".

An ENDOTHERMIC Process (viz. a process where energy is absorbed)
(∑ H for Products > ∑ H for Reactants, therefore ΔH is positive)



The symbol "∑" means "The sum total of ....", and the symbol "Δ" means "the difference or change in ....".

An EXHOTHERMIC Process (viz. a process where energy is released/given out)
(∑ H for Products < ∑ H for Reactants, therefore ΔH is negative)



The symbol "∑" means "The sum total of ....", and the symbol "Δ" means "the difference or change in ....".

OK, we must now look at something called "**Activation Energy**".

We saw just now that Methane and Oxygen will react together in an Exhothermic reaction, but we need to refine that statement slightly.

If Methane and Oxygen are mixed together no reaction will occur, and even if the mixture were left for one billion years no reaction would take place. In order for the two reactants to react, something called **Activation Energy** has to be supplied to the mixture – and then there will be not just a reaction but (if there are sufficient quantities of the two reactants) there will be an explosion.

Activation Energy is the minimum amount of energy that must be supplied to the reactants in a reaction in order for the reaction to commence. If you wanted to participate in a bicycle race you would first need a bicycle – and *Activation Energy* is the so-to-speak the "bicycle" that reactants need to participate in a reaction.

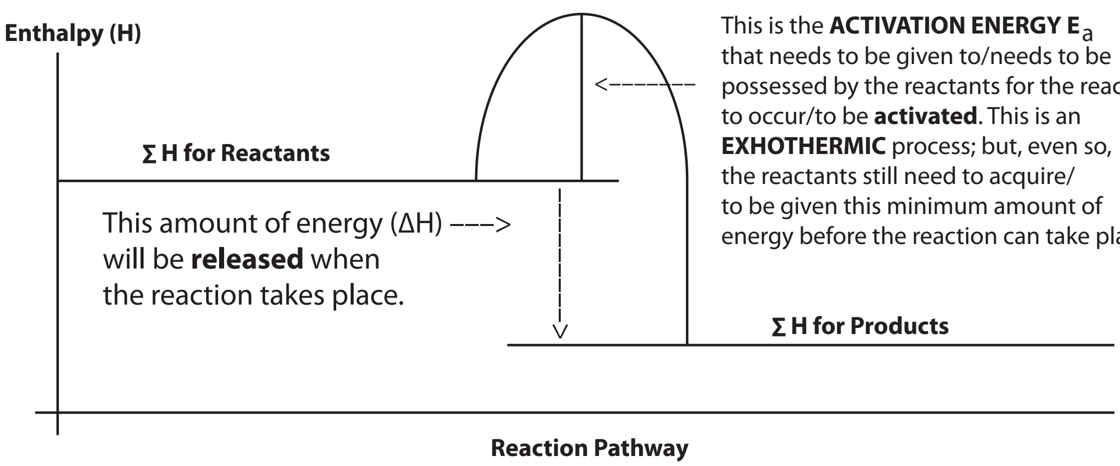
ACTIVATION ENERGY (E_a)

What is Activation Energy? Activation Energy is the minimum amount of energy that reactants must possess in order for a reaction to occur. If reactant molecules do not possess that *minimum*/that *threshold* amount of energy, then they will **not** react on collision with each other.[4]

For a reaction to occur the reactants must possess sufficient energy to break the existing bonds that need to be broken and then form the new bonds that need to be formed. It is a bit like a boy and a girl who are madly in love! If the girl is madly in love with a boy, then it would not matter how many times she collided with another boy, her love for her existing partner would be so great that the existing bond could not be broken – but if she only liked (as opposed to loved) her existing partner, then the new boy in her life could perhaps break the existing bond and then form a new bond with her!

The analogy above is the sort of analogy that you might find in a twelve year old girl's magazine – but even so it is a good analogy because it makes the point that for a reaction to occur between reactants, they have to collide **with *sufficient* force for existing bonds to be broken** and new bonds to form – otherwise the colliding reactants will merely bounce off each other without **reacting** with each other. **For a reaction to occur, the minimum amount of energy that reactants must possess when they collide with each other is called the ACTIVATION ENERGY, E_a**.

In an **Exothermic** process, energy is given off by the reaction – but, even so, it takes an extra amount of energy to START the reaction off in the first place! Just imagine the gas cooker in a home. At room temperature a match needs to be struck in order to start the gas burning – but, after that, the molecules of gas that are burning i.e. Methane and Oxygen (from the air) are reacting together and giving off a large amount of heat to cook the food, and this heat *also provides the Activation Energy required for further molecules of Methane and Oxygen to react / to combust (until the tap in the cooker is turned off)* . It is the burning match that provides the initial Activation Energy, E_a, for the *commencement* of the reaction of the Combustion of Methane to occur at Room Temperature – but after that the exothermic nature of the process provides all the activation energy necessary to continue the process!



Here energy has been released therefore ΔH will have a negative sign.

If you are ever asked why a given exothermic reaction does not occur spontaneously (at Room Temperature) – the answer that you must give is that "the reaction cannot occur until the reactants possess the appropriate amount of ***Activation Energy***.

An ENDOTHERMIC Process

i.e. one where ∑ H for the Products > ∑ H for the Reactants (therefore ΔH is positive).

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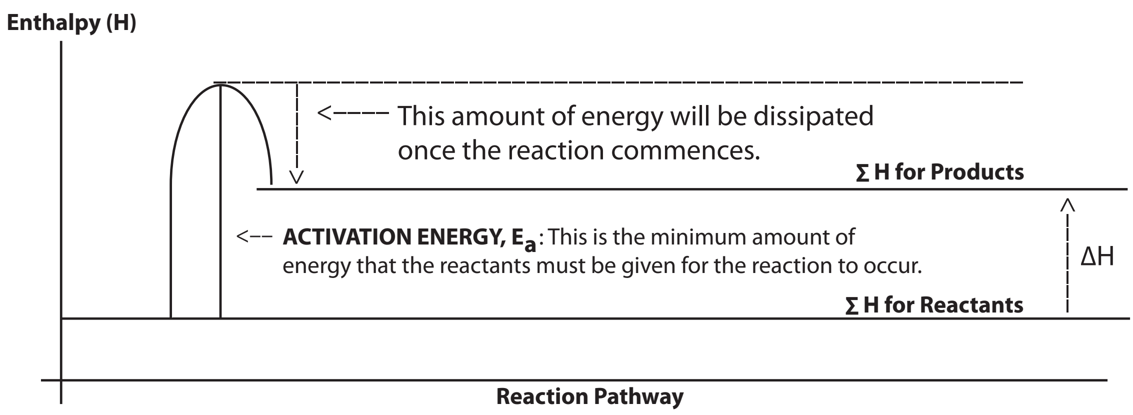
[4] Please note that for a reaction to take place, if more than one species is involved then (i) **the Reactant species must collide,** (ii) with the qualifying amount of Activation Energy (E_a), whereupon (iii) existing bonds must be broken, and (iv) new bonds must be formed – and unless all **four** things happen then **no reaction will take place**.



NB ΔH is the amount of energy absorbed/consumed in the reaction and it is given by ∑ H for the Products minus ∑ H for the Reactants. (Here energy has been sucked in/ consumed therefore ΔH will have a <u>positive</u> sign.)

In an Exothermic reaction, the energy constantly being released by the reaction allows the process to continue once it has been started; but, in an **Endothermic** reaction, the process has to **suck in energy** continuously from the surrounding atmosphere – *and by doing so, THAT is what enables it to keep going*! If you were to place the liquid reactants for an Endothermic reaction in a metal bowl and start the reaction going, then the metal bowl would become very cold to touch *because the reactants would be sucking heat out of the bowl (and thereby heat out of the surrounding air), and the bowl would then be sucking heat out of your hand (and this is why the bowl would feel cold)*.

In exams, you will always be questioned on two things (a) the theory involved, and (b) the ability to apply that theory. In questions on Enthalpy changes, when the examiners ask you how to calculate a ΔH for a particular experiment – and in this Section you will therefore be required to learn **FOUR** ways to calculate a ΔH viz.

1) using experimental data where you use the formula **"ΔH = m . c . ΔT"** (Section 7.1, page 113)
2) using Enthalpy Profile Diagrams (pp 116-17)
3) using Hess' Cycles, and
4) using Average Bond Enthalpies/Bond Energy Terms (Section 7.4, pp118-19).