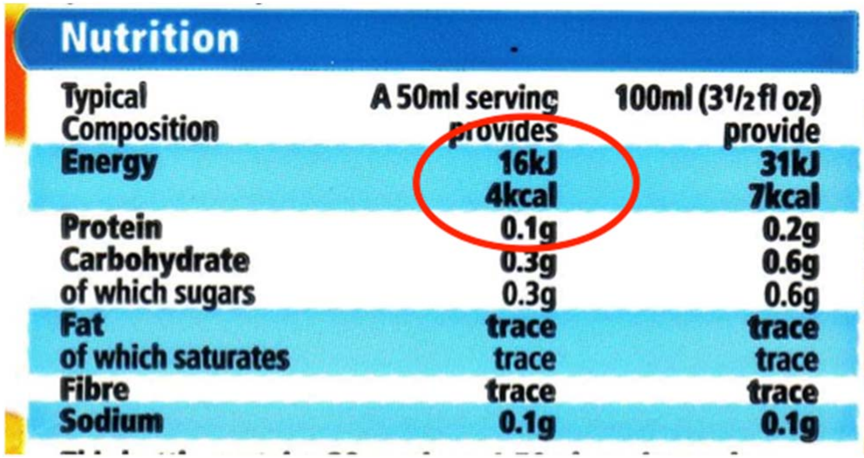
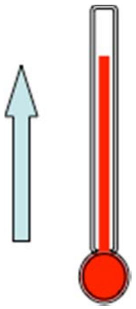
Enthalpy Changes

All substances contain chemical energy, called **enthalpy**. Like any kind of energy it is measured in Joules (previously energy was measured in Calories). When reactions happen, energy is given out or taken in – these are **enthalpy changes**.



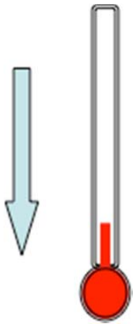
Nutrition		
Typical Composition	A 50ml serving provides	100ml (3½fl oz) provide
Energy	16kJ 4kcal	31kJ 7kcal
Protein	0.1g	0.2g
Carbohydrate	0.3g	0.6g
of which sugars	0.3g	0.6g
Fat	trace	trace
of which saturates	trace	trace
Fibre	trace	trace
Sodium	0.1g	0.1g

Note: 1 calorie = 4.2 Joules



In an EXOTHERMIC reaction:
Chemical energy (enthalpy) is being turned into heat energy which is transferred to the surroundings, so the temperature we measure **increases**.

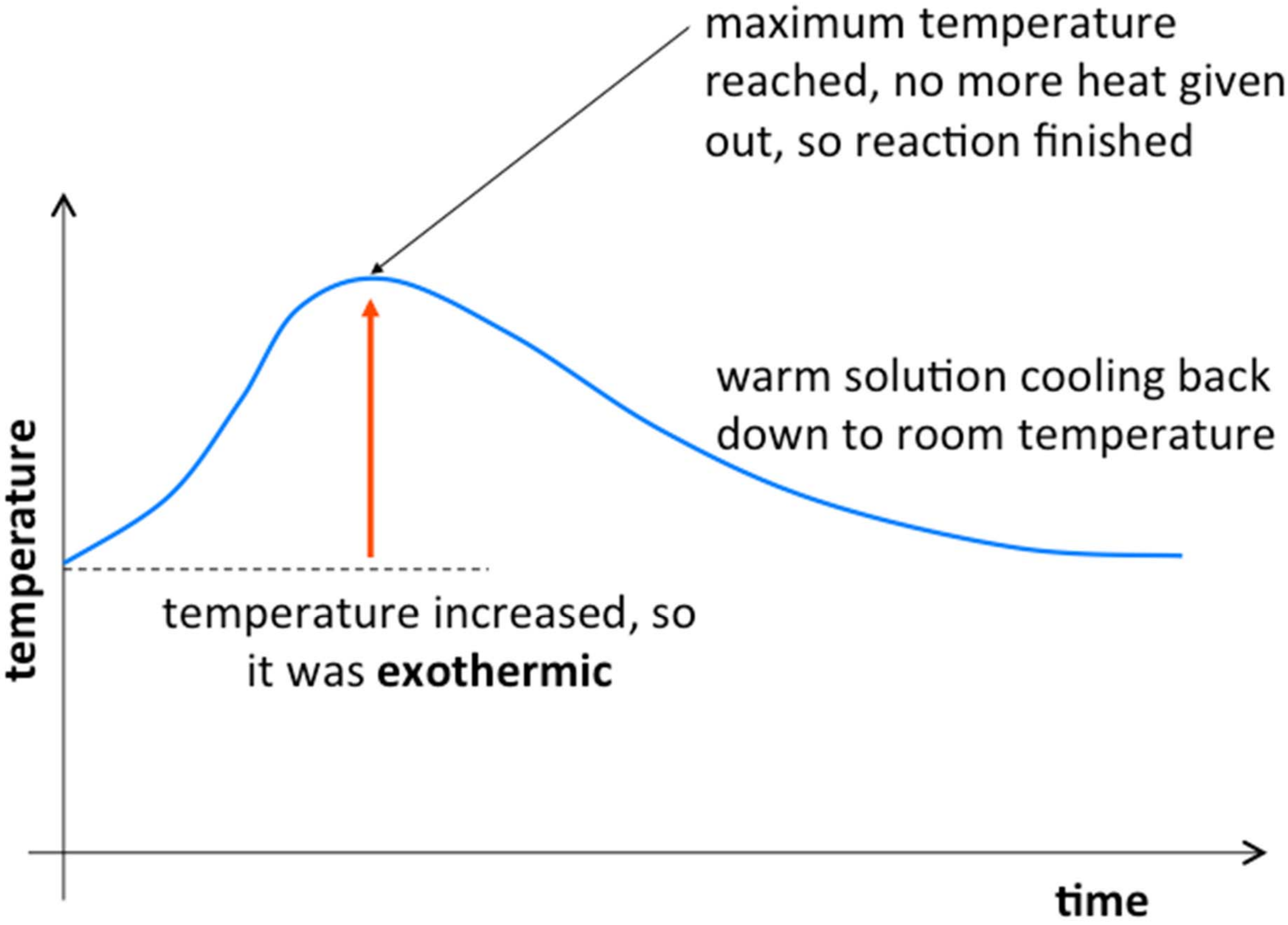
- combustion of fuels (including respiration)
- many oxidation reactions
- neutralisations



In an ENDOTHERMIC reaction:
Heat energy is taken from the surroundings and converted into chemical energy (enthalpy), so the temperature **decreases**, or we have to heat the reaction constantly to make it work.

- thermal decompositions
- photosynthesis (light energy in !)

Temperature change during an exothermic reaction



Uses of exothermic reactions:

REUSABLE

DISPOSABLE

Hand Warmers

Hand warmers work by **exothermic reaction**; the reusable one is 'recharged' using the reverse **endothermic reaction**.

Self-heating cans

When the seal is broken, an **exothermic reaction** is used to heat the contents of the can, ready for eating/drinking.

Uses of endothermic reactions:

Cold packs

Non-refrigerated sports injury cold packs use an **endothermic reaction** to cool the pack down quickly ready to be used on the injury to reduce swelling.

Calorimetry

Calorimetry is the technique of measuring **how much energy** a food or fuel releases when it burns. Foods also release energy when they are used for respiration - this is also an exothermic reaction, exactly the same as burning, but taking place inside the body.

We calculate the heat energy released during a reaction, Q, by measuring the temperature change when a known mass of water is heated using this energy.

$$Q = m\ c\,\Delta T$$

*c is called the **Specific Heat Capacity**,*
and its value will be given in the question

Q is energy released, in Joules
m is the mass of water (in g)
ΔT is the change in temp. (in ˚C)
c is a constant = 4.2 (J/g/˚C)

Working out energy released:

e.g. In a calorimetry experiment, a fuel sample was burned and produced a temperature change of 25°C in the calorimeter, which contained 100g of water. **c = 4.2 J/g/˚C**

$$Q = m \times c \times \Delta T \qquad = 100 \times 4.2 \times 25$$

$$= \underline{10{,}500\ J\ \ (\text{or } 10.5\ kJ)}$$

Comparing enthalpy changes

To <u>compare</u> reactions, we can calculate the energy released <u>per mol</u>e of reactant used, which we call the **molar enthalpy change**, ΔH.

To do this we work out the energy released as before, then divide by the moles reactant used.

$$\Delta H = -Q\ /\ moles$$

Example:

A 0.5g sample of ethanol is burnt, raising the temperature of 250g of water from 10˚C to 28˚C. When 0.8g of butane is burnt, the temperature of 250g of water increases from 10˚C to 40˚C. Which fuel produces the most heat energy per mole of fuel burnt ? **c = 4.2 J/g/˚C**

For ethanol:

Q = (4.2 x 250 x 18) = 18,900 J

Mol ethanol = mass/Mr = 0.5/46 = 0.01087

ΔH = -Q/moles = -18,900/0.01087

= -1,738,730 J/mol

For butane:

Q = (4.2 x 250 x 30) = 31500 J

Mol butane = mass/M_r = 0.8/58 = 0.01379

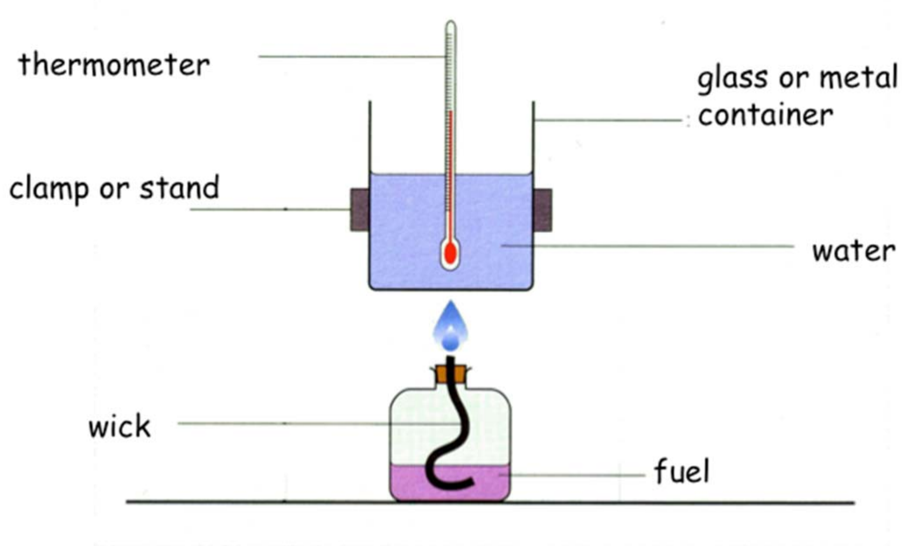
ΔH = -Q/moles = -31500/0.01379

= -2,284,264 J/mol

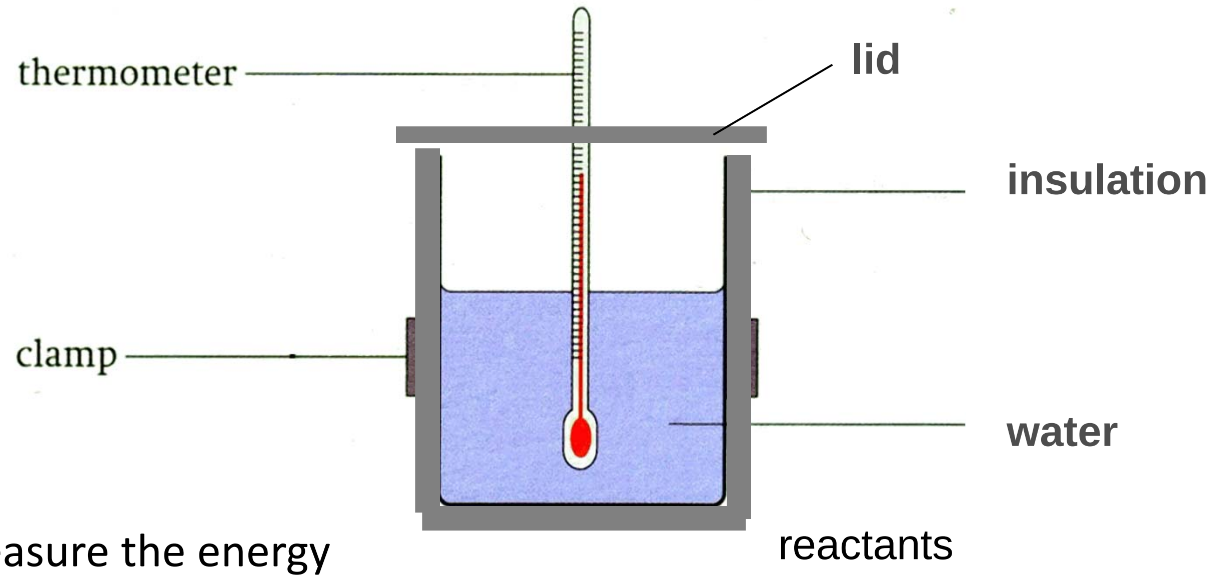
A simple calorimeter not very accurate, but can be used to **compare** the amount of energy released by different fuels or foods.

It is not accurate, or precise, as there are several major sources of experimental error:

- Heat is lost around the sides of the calorimeter, so not heating the water
- Incomplete combustion of the fuel may release less energy
- The calorimeter is not insulated so heat is lost from the water



Temperature changes in other reactions



Calorimetry can be used to measure the energy changes in:

- reactions of solids with solutions (or water)
- neutralisation reactions   etc.

The amount of energy produced by a chemical reaction in solution can be found by mixing the reagents in an **insulated container** (e.g. polystyrene cup) and measuring the temperature change caused by the reaction.

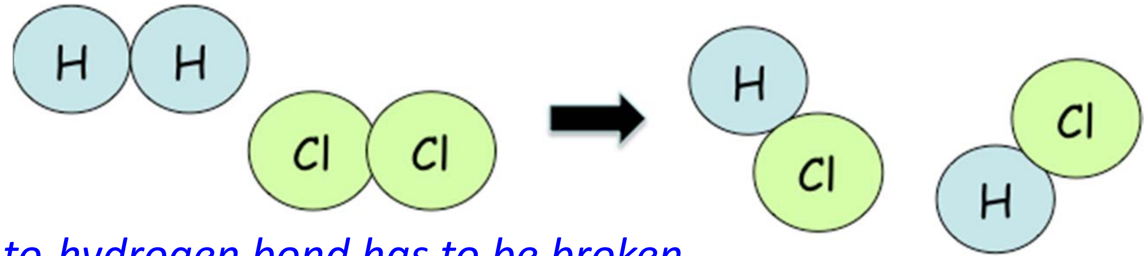
The total mass of the reactants is used in the calculation of $Q = m\, c\, \Delta T$

The amount of each reactant (in moles) needs to be checked against the balanced chemical equation to see which reactant is NOT in excess. Moles of the reactant NOT in excess should be used in the calculation of $\Delta H = -Q / \text{moles}$

What happens during a reaction
We know that when a reaction takes place, bonds are broken, and new bonds formed.

Consider the reaction between hydrogen and chlorine molecules to make hydrogen chloride, and think about what bonds have to be broken, and what bonds formed:

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



One hydrogen-to-hydrogen bond has to be broken
One chlorine-to-chlorine bond has to be broken
Two hydrogen-to-chlorine bonds are formed

Energy is needed to break bonds. That's why molecules have to collide with sufficient energy for a successful reaction to take place. **Activation energy** is the energy required to break all the necessary bonds in the reactants.

The **stronger** the bonds are, the **more energy** is needed to break them (bond energy).

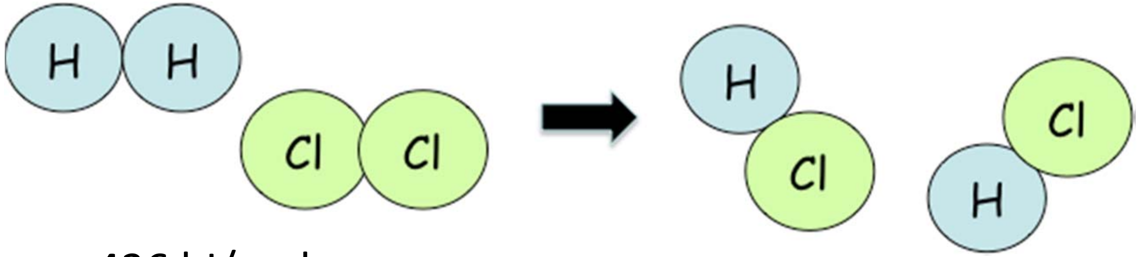
The table shows how much energy different types of bonds require to break: the **average bond energy**.

bond	average bond energy in kJ/mol
H-H	436
C-H	412
C-C	348
Cl-Cl	242
O=O	496
O-H	463
H-Cl	431
C=O	743
N≡N	944
H-Br	366
Br-Br	193

So bond breaking requires heat energy to be taken from the surroundings and used to break the bonds. The surroundings get cooler. Bond breaking is **ENDOTHERMIC**

When new bonds form, energy is given out. This causes the surroundings to heat up. Bond forming is **EXOTHERMIC.** (The amount of energy given out is equal to the bond energy for that bond)

Consider our reaction between hydrogen and chlorine again:



One H-H bond is broken:		436 kJ/mol
One Cl-Cl bond is broken:		242 kJ/mol
TOTAL ENERGY TAKEN IN	=	**678 kJ/mol**
Two H-Cl bonds are made		2 x 431 kJ/mol
TOTAL ENERGY GIVEN OUT	=	**862 kJ/mol**

Overall, is less energy is taken in than is given out. This means overall energy is being released – this reaction is exothermic.

The molar enthalpy change for reactions can be calculated:

ΔH = energy taken in – energy given out

In this example the molar enthalpy change is **-184 KJ/mol**. We see that the value of ΔH is negative for all exothermic reactions.

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H-H	436
C-H	412
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Cl-Cl	242
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O-H	463
H-Cl	431
C=O	743
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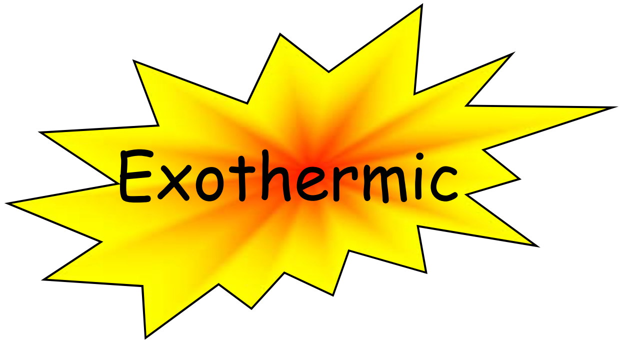
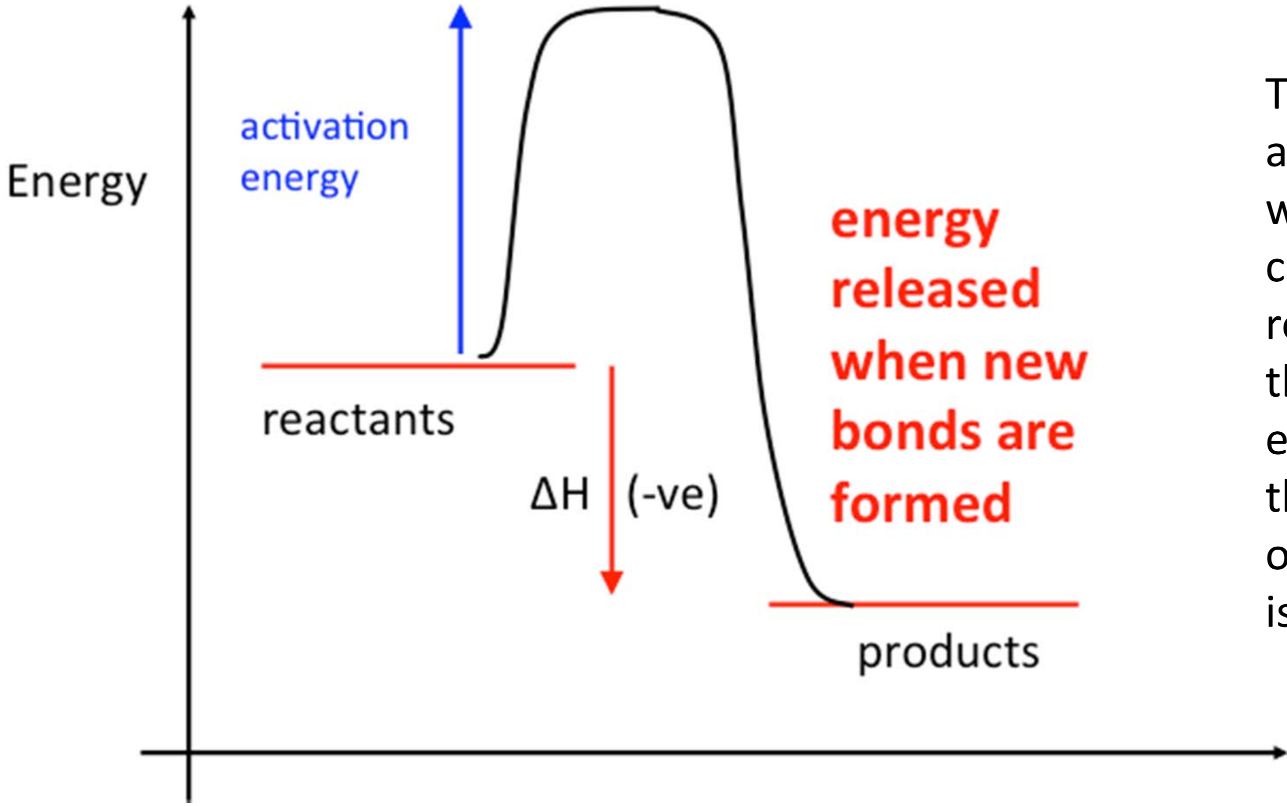
H-Br	366
Br-Br	193

In an **exothermic reaction, ΔH is always negative** (the energy of the products is lower than that of the reactants).

Explanation: <u>Less</u> energy is <u>used</u> to break the bonds in the reactants than is <u>released</u> when the bonds in the products are formed.

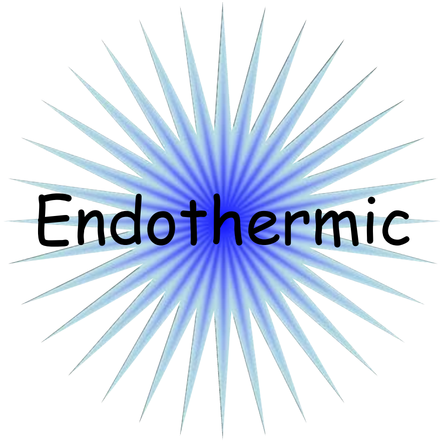


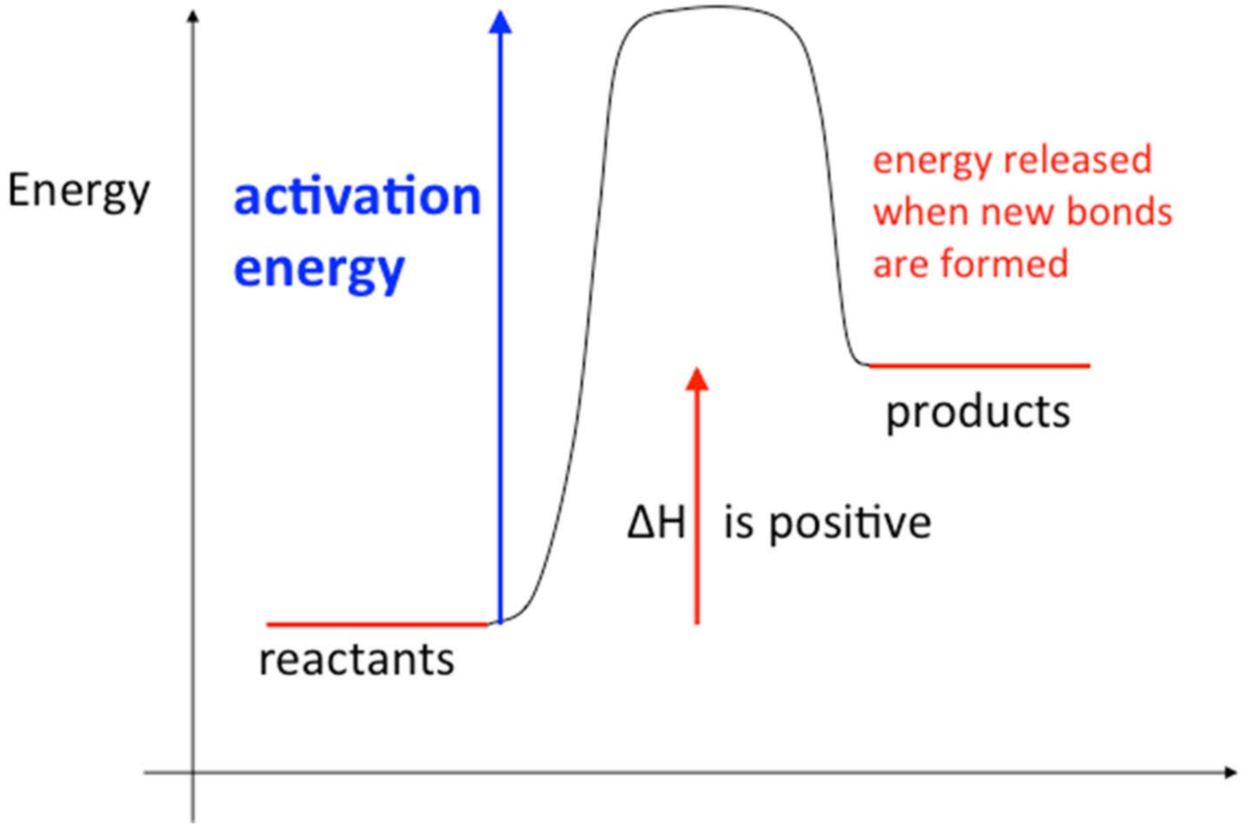


This can be illustrated using an **<u>energy level diagram,</u>** which shows the energy changes during the reaction. We can see that the products are of lower energy (more stable) than the reactants, and that the overall energy change (ΔH) is negative.

In an **endothermic reaction, ΔH is always positive** (the energy of the products is higher than that of the reactants)

Explanation: <u>More</u> energy is <u>used</u> to break the bonds in the reactants than is <u>released</u> when the bonds in the products are formed.





This can be illustrated using an **<u>energy level diagram,</u>** which shows the energy changes during the reaction. We can see that the products are at higher energy (less stable) than the reactants, and that the overall energy change (ΔH) is positive.

Effect of a catalyst

A catalyst increases the rate of a reaction, without being used up itself.

It does this by **lowering the activation energy** for the reaction, by providing an **alternative pathway**. On the catalyst surface the reactants have their bonds weakened, so that less energy is needed to break these bonds.

This can also be illustrated using an **energy level diagram**.

The effect of the catalyst is to make the activation energy smaller, but less energy is released when the bonds in the products are formed, so overall ΔH is the same. The energy released during a reaction is not affected by the use of a catalyst.