Exothermic reactions

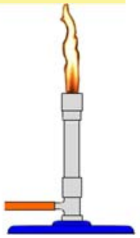
When a chemical reaction occurs, energy is transferred to or from the surroundings - and there is often a temperature change.

Exothermic reactions **transfer energy to the surroundings**. The energy is usually transferred as heat energy, causing the reaction mixture and its surroundings to become hotter. The temperature increase can be detected using a thermometer. Some examples of exothermic reactions are:

Combustion (burning)

Many oxidation reactions, for example rusting

Neutralisation reactions between acids and alkalis



When a flame burns it transfers heat to its surroundings.

Exothermic reactions can be used for everyday purposes. For example, hand warmers and self-heating cans for drinks (such as coffee) use exothermic reactions.

Endothermic reactions

These are reactions that **take in energy from the surroundings**. The energy is usually transferred as heat energy, causing the reaction mixture and its surroundings to get colder. The temperature decrease can also be detected using a thermometer.

Some examples of endothermic reactions are:

Electrolysis

The reaction between ethanoic acid and sodium carbonate

The *thermal decomposition* of calcium carbonate in a blast furnace

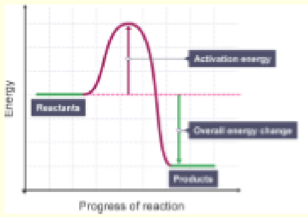
Endothermic reactions can be used for everyday purposes. For example, certain sports injury cold packs use endothermic reactions.

1. Exothermic & Endothermic reactions

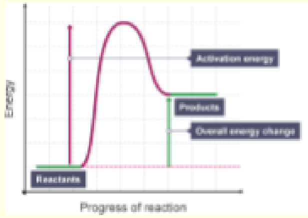
(a) **Exo**thermic: a reaction that **gives out** energy to the surroundings (temperature **increases**)

(b) **Endo**thermic: a reaction that **takes in** energy from the surroundings (temperature **decreases**)

(c) **Energy level diagram** (reaction profile) shows the energy changes taking place in a chemical reaction



(i) Exothermic reaction



(ii) Endothermic reaction

(e) **Activation energy** the minimum amount of energy required for a reaction to take place

(f) In a reaction, chemical bonds are **broken** and **made**.

Breaking bonds = **require** energy = **endo**thermic process
Making bonds = **release** energy = **exo**thermic process

More energy required than released = endothermic reaction
More energy released than required = exothermic reaction

Electrode reactions

There are two *electrodes* in a hydrogen-oxygen *fuel cell*.

You should be able to write the **balanced equations** for the reactions when you are given the formulae of the *ions* and *products*.

For example:

At the cathode (negative electrode): $H_2(g) - 2e^- \rightarrow 2H^+(aq)$

At the anode (positive electrode): $4H^+(aq) + O_2(g) + 4e^- \rightarrow 2H_2O(g)$

The reaction at the cathode is an *oxidation reaction* because hydrogen loses *electrons*, and the reaction at the anode is a *reduction reaction* because hydrogen ions gain electrons. The overall reaction in the fuel cell is a *redox reaction*.

You could try to remember the following to help jog your memory:

Bond energy calculations

You can calculate the energy change in a reaction using average bond energies. **Bond energy** is the amount of energy needed to break one mole of a particular bond.

To calculate bond energy

Add together the bond energies for all the bonds in the reactants – this is the 'energy in'.

Add together the bond energies for all the bonds in the products – this is the 'energy out'.

Calculate the energy change = energy in – energy out.

Worked example – an exothermic reaction

Hydrogen and chlorine react to form hydrogen chloride gas:

H-H + Cl-Cl → 2 × (H-Cl)

Bond	Bo
H-H	436
Cl-Cl	243
H-Cl	432

Energy in = 436 + 243 = 679 kJ/mol

Energy out = 2 × 432 = 864 kJ/mol

Energy change = in – out = 679 – 864 = **–185 kJ/mol**

The energy change is negative, showing that energy is released to the surroundings in an exothermic **reaction**.

2. Cells and Batteries (triple only)

(a) **Cell** (electrical): contains chemicals that react together to release energy.

The voltage of a cell can be changed by changing the *type of electrode* and the *type of electrolyte*

(i) **Electrode**: An electrical conductor used in a cell

(ii) **Electrolyte**: A solution or molten substance that is broken down during electrolysis.

(b) **Battery**: Multiple cells connected together.

(i) **Non-rechargeable battery** – stops producing electricity when one of the reactants has been used up!

(ii) **Rechargeable battery** – the chemical reaction can be reversed using an external electrical current so the battery can be reused.

(c)**Fuel Cell**: Efficient way of producing electrical energy where a fuel is oxidised electrochemically to produced a potential difference (or voltage) (normally *hydrogen*).

Anode: $2H_2 \rightarrow 4H^+ + 4e^-$
Cathode: $O_2 + 4H^+ + 4e^-$ $2H_2O$

+ less stages, less polluting, more efficient