Exothermic and endothermic reactions

Exothermic reactions **release** thermal energy (heat) into their surroundings. They can occur spontaneously and some are explosive. Most chemical reactions are exothermic. Temperature **increases**.

What are some examples?

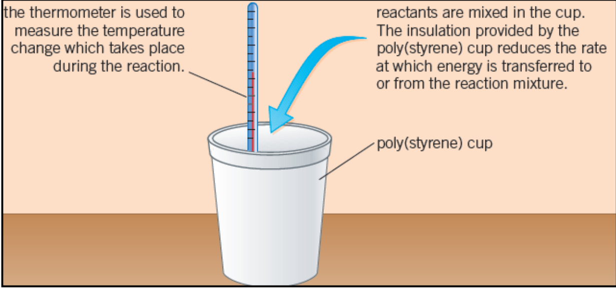
- combustion
- respiration
- neutralization of acids with alkalis
- reactions of metals with acids
- Mg (s) + HCl (aq) → $MgCl_2$ (aq) + H_2 (g)
- the Thermite Process.
- Endothermic reactions absorb thermal energy, and so cause a **decrease** in temperature.

What are some examples?

- thermal decomposition, e.g. calcium carbonate in a blast furnace
- photosynthesis
- some types of electrolysis
- Sherbet
- NH_4NO_3 (s) + H_2O (l) → NH_4^+ (aq) + NO_3^- (aq)

Investigating temperature changes

Record the initial temperatures of any solutions, and the maximum and minimum temperatures reached in the course of the reaction.



Using energy transfers from reactions

- Exothermic changes can be used in hand warmers and self heating cans. Crystallisation of the supersaturated solution is used in reusable warmers. However, disposable, one-off hand warmers heat up the surrounding for longer.
- Endothermic changes can be used in instant cold packs for sports injuries.

Reaction profiles and Activation energy

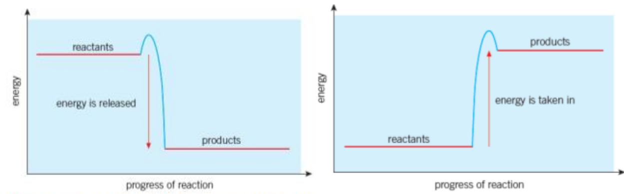


Figure 1 *The reaction profile for an exothermic reaction* **Figure 2** *The reaction profile for an endothermic reaction*

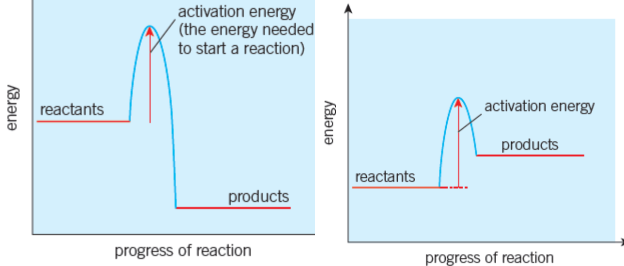


Figure 3 *This reaction profile shows the activation energy for an exothermic reaction*

Figure 4 *This reaction profile shows the activation energy for an endothermic reaction*

Bond breaking is endothermic whereas bond making is exothermic.

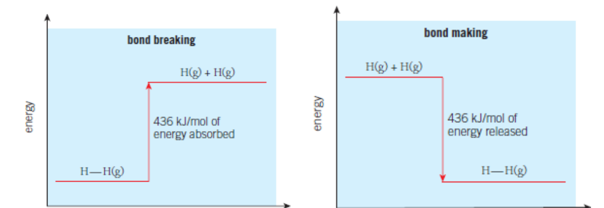
Bond energy calculations

The energy needed to break a bond between 2 atoms is called the **bond energy** for that bond. They are measured in KJ/mol.

Table 1 *Common bond energies*

Bond	Bond energy in kJ/mol
C—C	347
C—O	358
C—H	413
C—N	286
C—Cl	346
Cl—Cl	243

Bond	Bond energy in kJ/mol
H—Cl	432
H—O	464
H—N	391
H—H	436
O═O	498
N≡N	945

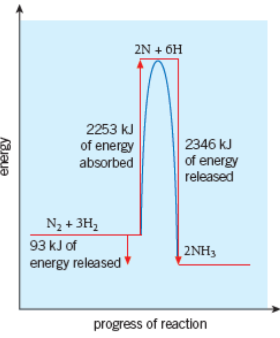


Breaking and making a particular bond always involves the same amount of energy

The formation of ammonia. The energy released, 93KJ, is from the formation of 2 moles of ammonia (see balanced equation below). So if you wanted to know the energy change for the reaction per mole of ammonia formed, it would release exactly half this, i.e. 46.5kJ/mol.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

- In chemical reactions, energy must be supplied to break the bonds between atoms in the reactants.
- When new bonds are formed between atoms in a chemical reaction, energy is released.
- In an exothermic reaction, the energy released when new bonds are formed is greater than the energy absorbed when bonds are broken.
- In an endothermic reaction, the energy released when new bonds are formed is less than the energy absorbed when bonds are broken.
- You can calculate the overall energy change in a chemical reaction using bond energies.



Cells and batteries

$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

The sulfate ions do not change in the displacement reaction above. They are spectator ions.

So you can leave them out of the equation and write an ionic equation:

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

You can think of this redox reaction as two half equations.

One will represent reduction:

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

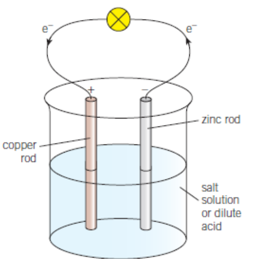
The Cu^{2+} ions are reduced to Cu.

The other will be an oxidation reaction:

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$$

The Zn atoms are oxidised to Zn^{2+} ions.

An electrical cell made from zinc and copper → The electrons flow from the more reactive metal (zinc) to the less reactive metal (copper). So zinc acts as the negative terminal of the cell, providing electrons to the external circuit.

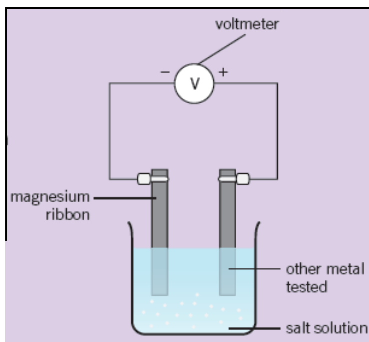


Cells and batteries continued…

- Metals lose electrons and form positive ions.
- When 2 metals are dipped in a salt solution and joined by a wire, the more reactive metal will donate electrons to the less reactive metal. This forms a simple electrical cell.
- The greater the difference in reactivity between the 2 metals, the higher the voltage produced by the cell.

Investigating chemical cells

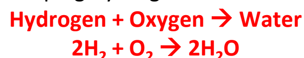
This apparatus is used to investigate the voltage produced by different metals paired with magnesium ribbon. You can compare magnesium against zinc, iron, copper & tin in your electrical cells.



Fuel Cells

Scientists are developing hydrogen as a fuel.

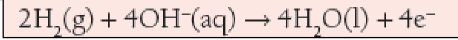
Hydrogen + Oxygen → Water



$2H_2 + O_2 \rightarrow 2H_2O$

- The world relies on fossil fuels. However, they are non-renewable and they cause pollution.
- Hydrogen is one alternative fuel. It can be burned in combustion engines or used in fuel cells to power vehicles.
- Hydrogen gas is oxidised and provides a source of electrons in the hydrogen fuel cell, in which the only waste product is water.

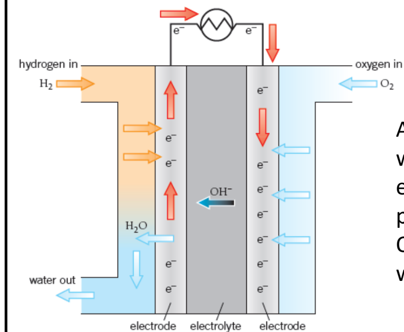
Hydrogen gas is supplied as a fuel to the negative electrode. It diffuses through the graphite electrode and reacts with hydroxide ions to form water and provides a source of electrons to an external circuit.



$$2H_2(g) + 4OH^-(aq) \rightarrow 4H_2O(l) + 4e^-$$

Oxygen is supplied to the positive electrode. It diffuses through the graphite and reacts to form hydroxide ions, accepting electrons from the external circuit.

$$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$$



A hydrogen fuel cell which has an alkaline electrolyte, such as potassium hydroxide. Only waste product is water.

Advantages of hydrogen fuel cells –

1) Do not need to be electrically recharged
2) No pollutants are produced
3) Can be a range of sizes for different uses

Disadvantages of hydrogen fuel cells–

1) Hydrogen is highly flammable
2) Hydrogen is sometimes produced for the cell by non-renewable sources
3) Hydrogen is difficult to store