Energy Changes

CHEMISTRY • REACTIONS • ENERGY CHANGES

Section 1: Energy Changes

• What are exothermic reactions and endothermic reactions?

All chemical substances contain energy, and all chemical reactions involve some kind of energy change. If heat energy is given out during a reaction we say it is **exothermic**, whereas, if heat energy is taken in during a reaction we say it is **endothermic**.

The heat content of chemical substances (measured at constant pressure) is called the enthalpy (H), which is measured in kilojoules per mol ($kJmol^{-1}$). We cannot measure enthalpy directly, but we can measure the enthalpy change (ΔH) during a chemical reaction, which is defined as:

$$\Delta H = \text{(total enthalpy of products)} - \text{(total enthalpy of reactants)}$$

Burning carbon to make carbon dioxide is an exothermic reaction and the enthalpy change (ΔH) is negative, because the total enthalpy of products is less than the total enthalpy of reactants:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$
$$\Delta H = -393.5\ kJmol^{-1}$$

If sodium hydrogencarbonate (baking soda) is used to make a cake there is a thermal decomposition, resulting in bubbles of carbon dioxide gas which cause the cake to rise. The reaction is endothermic and the enthalpy change (ΔH) is positive, because the total enthalpy of products is more than the total enthalpy of reactants:

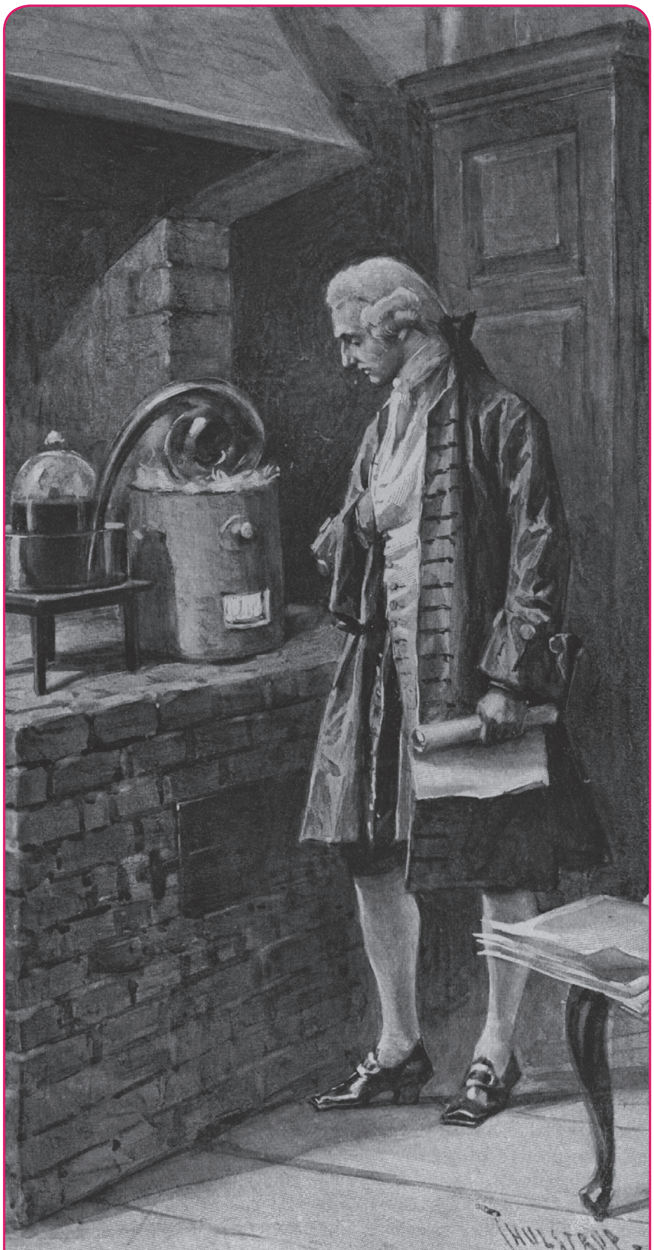
$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$$
$$\Delta H = +91.6\ kJmol^{-1}$$

• Suggested Films

- Oxygen and Combustion
- Energy Changes of Reactions
- The Elements: Oxygen

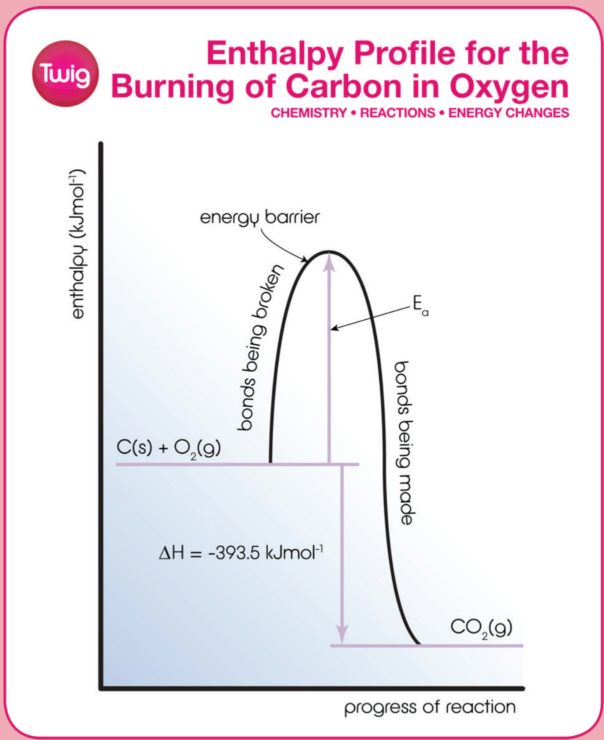
• Worksheet Questions

- Questions 1 and 2



An artist's impression of the French chemist Antoine Lavoisier (1743 - 1794) at work

DIAGRAM 01:



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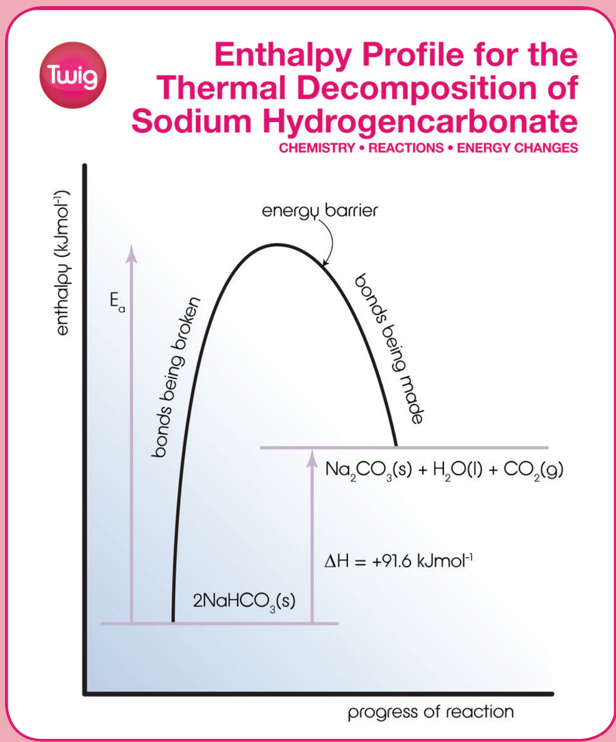


• What is the connection between bonds and energy changes during reactions?

All chemical reactions involve bonds being broken and new bonds being made. Bond **breaking** is always **endothermic**, because heat energy has to be taken in to pull the atoms apart. Bond **making** is always **exothermic,** because heat energy is given out when the bonds are made. For example, if we break a hydrogen molecule into its gaseous atoms:

$$H\text{-}H(g) \rightarrow H(g) + H(g)$$
$$\Delta H = +436\ kJmol^{-1}$$

During a chemical reaction involving substances with covalent bonds, the enthalpy change (ΔH) can be defined as:

(sum of enthalpy changes when bonds are broken) - (sum of enthalpy changes when bonds are made)

In the following reaction where hydrogen gas and chlorine gas react to make hydrogen chloride gas:

$$H_2(g) + C_{l2}(g) \rightarrow 2HCl(g)$$

We can show the covalent bonds involved as:

$$(H\text{-}H) + (Cl\text{-}Cl) \rightarrow 2(H\text{-}Cl)$$

Fireworks involve exothermic reactions

The enthalpy change in breaking the H-H and Cl-Cl bonds is 436 + 243 = +679 $kJmol^{-1}$
The enthalpy change involved in making the H-Cl bonds is -2 x 432 = -864 $kJmol^{-1}$
(The minus sign is there because heat is given OUT when bonds are formed; we need ‘2’ because there are 2 moles of HCl.)

The overall enthalpy change for this reaction
ΔH = +679 – 864 = -185 $kJmol^{-1}$, so the reaction is exothermic.

• Suggested Films

- Oxygen and Combustion
- Energy Changes of Reactions
- Covalent Bonding

• Are some reactions always exothermic and some reactions always endothermic?

Combustion is always exothermic, which is why we burn fuels such as methane to release heat and provide energy which we can use to cook food or heat our houses.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$
$$\Delta H = -890.3\ kJmol^{-1}$$

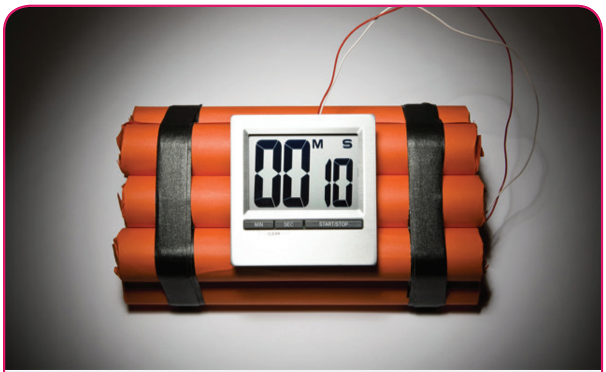
Explosions, such as those in fireworks and in the use of dynamite for quarrying, always involve exothermic reactions. Explosions are very rapid combustion reactions, usually involving the simultaneous evolution of gas. The pressure wave made by the rapidly expanding hot gas produces the sound of the explosion.

Aerobic respiration in cells is also exothermic, but, unlike in an explosion, the energy released comes out in a much more controlled and less destructive way. It can be used to do useful tasks such as eating, walking and thinking! Glucose molecules react with oxygen, making carbon dioxide and water and releasing a large amount of energy.

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$
$$\Delta H = -2870\ kJmol^{-1}$$

The reactions involved in cooking and digesting food are often endothermic. They involve breaking bonds in large, complex molecules, such as proteins, fats and carbohydrates, creating simpler, smaller molecules. Thermal decomposition reactions, such as the cracking of hydrocarbons and the thermal decomposition of baking soda, are also usually endothermic, as is the thermal decomposition of limestone (calcium carbonate), which is broken down into calcium oxide and carbon dioxide in a lime kiln:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$
$$\Delta H = +178.3\ kJmol^{-1}$$



Dynamite was invented by Alfred Nobel in 1867

• Suggested Films

- Oxygen and Combustion
- The Elements: Oxygen
- Energy Changes of Reactions
- Nobel and Dynamite
- How Do Fireworks Work?
- The Hindenburg Disaster
- Oxidation Reactions

Section 2: Rates of Reaction

• Why are some reactions fast and other reactions slow?

Some reactions, such as explosions, are over in a fraction of a second. Others, like rusting, corrosion, ripening fruit, fermentation and digestion are rather slower. Yet others, such as fossilisation, the formation of limestone and sandstone, the formation of crude oil, natural gas and coal may take millions of years. We measure the speed of a reaction by its rate, which we can define as either the amount of product made per second or the amount of reactant used up per second. By 'amount' we mean 'mass' or 'volume' depending on the substances involved.

To understand why there is such a huge range of rates of reaction we have to understand how reactions take place. For a reaction between substance **A** and substance **B** to take place, a particle of substance **A** has to collide with a particle of substance **B**. This is known as the 'collision theory' of reactions. However, merely colliding is not enough: they have to collide with a certain minimum energy before the reaction can take place. This is known as the **activation energy**, which is a threshold value, and is symbolised as E_a. Activation energy is a threshold value because some of the bonds inside the particles have to break before the reaction can take place. Unless these bonds break, there will be no reaction. The activation energy is therefore usually endothermic, as energy is taken in order to break the bonds. Unless the particles have energy equal to or greater than E_a they will not react, no matter how often they collide.

Diagrams 1 and 2 show the 'hill' between the reactants and products, known as the **energy barrier**. The upward slope of the 'hill' is due to energy being needed to break bonds, and the downward slope of the 'hill' is due to energy being released when new bonds are made. The activation energy is the height of the 'hill', measured from the level of the reactants.

Very slow reactions have very high activation energies. For example, the reactions involved in fossilisation are reactions between solids, where the bonds between the particles are very strong and require large amounts of energy to break. This makes the fossilisation process extremely slow. However, even highly exothermic reactions like the combustion of carbon in oxygen can have high activation energies.

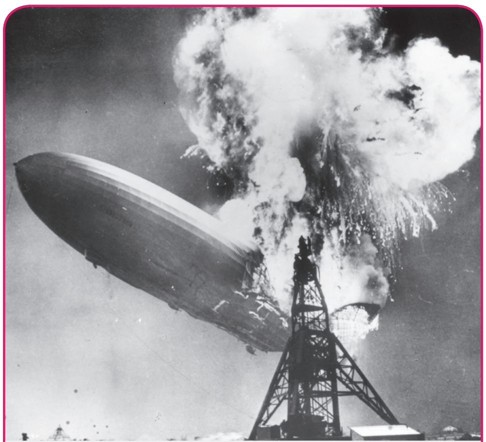
• Suggested Films

- Rates of Reaction: Basic
- Collision Theory
- Covalent Bonding
- Solids, Liquids and Gases
- Fermentation
- Ripening Fruit

• Worksheet Question

- Question 3

• What factors affect the rate of a reaction?


The Hindenburg airship was filled with the highly flammable gas hydrogen

Temperature is one of the most important factors. When cooking meat, too low a temperature means it stays raw; too high a temperature will make it burn. By raising the temperature we make the particles of the reactants move faster, so they collide more often per second. However, even more importantly, a higher proportion of the particles have energy which is greater than the activation energy, meaning that there are more successful collisions per second. Raising the temperature has a dramatic effect on the rate of the reaction, as a 10°C rise in temperature may double the rate, so a 60°C rise may cause a 64 fold increase in the rate. The effect is even more dramatic if the reaction is exothermic, as the heat released by the reaction raises the temperature even higher, making the rate accelerate. This "positive feedback" process may even become explosive. We can see this effect in many combustion reactions and in fireworks. The fire that destroyed the Hindenburg airship spread extremely rapidly, because once the reaction started, the highly exothermic combustion of hydrogen gas generated so much heat that the reaction became explosively violent.

The **surface area** of solids is another factor affecting the rate of reaction. When we cook vegetables we often chop them up into small pieces, as this makes them cook much faster due to an increased surface area. Finely powdered solids can be hazardous for this reason, as once they are ignited they can burn explosively.

The **concentration of solutions** is also important. Concentrated acids react faster than dilute acids: in a concentrated solution there are more particles per cubic centimetre, which means there will be more collisions per second.

• Suggested Films

- Oxygen and Combustion
- Energy Change of Reactions
- The Hindenburg Disaster
- The Elements: Oxygen
- Nobel and Dynamite
- How Do Fireworks Work?
- Oxidation Reactions

Pressure affects the rate of reaction in gases. A higher pressure forces particles closer together, which means particles collide more often per second, resulting in more successful collisions per second and, hence, a faster rate.

Finally, **catalysts** play a vital role is speeding up reactions, including the reactions in our bodies which keep us alive.

• What is a catalyst?

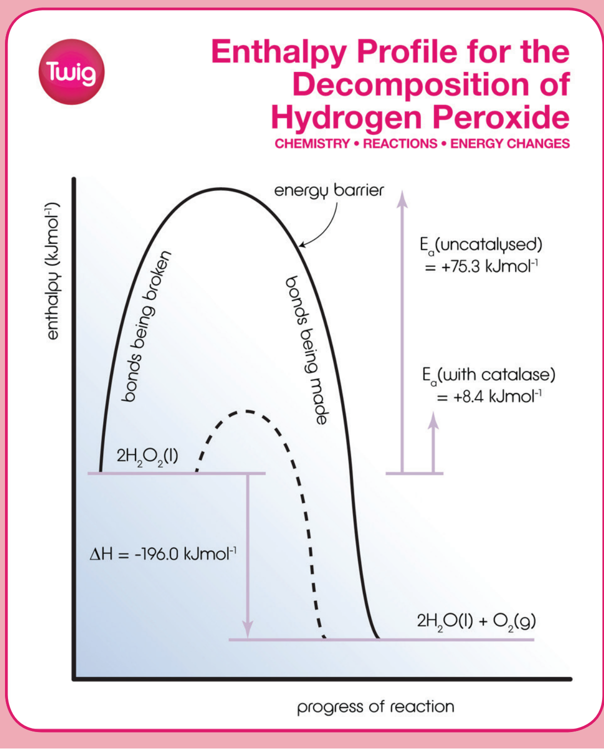
A catalyst is a substance that can speed up a reaction but which remains chemically unchanged at the end of the reaction. Catalysts are able to do this by providing an alternative route for the reaction with a lower activation energy. Industrial catalysts are often transition metals or their compounds.

Catalysts are very important for industrial chemistry as they speed up the production process enormously. They also allow the reaction to take place at a much lower temperature, thus reducing energy costs. However, catalysts cannot increase the yield of a reaction, they can only increase the rate at which the product is made.

An extremely important group of catalysts are **enzymes** or biological catalysts. Enzymes are protein molecules whose shape allows them to catalyse specific reactions with amazing efficiency. For example, hydrogen peroxide is produced in the body but is highly toxic to our cells. It can be removed by the highly exothermic (but very slow at room temperature) reaction:

$$2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$$
$$\Delta H = -196.0\ kJmol^{-1}$$
$$E_a = +75.3\ kJmol^{-1}$$

DIAGRAM 03:



Human tissues often contain catalase enzyme, and it is especially concentrated in the liver. In the presence of this enzyme, the activation energy for the decomposition of hydrogen peroxide drops dramatically to $E_a = +8.4\ kJmol^{-1}$, and the reaction rate speeds up by a factor of 10^{11} or 100,000,000,000. The rate of decomposition is 100 billion times faster than the reaction without catalase, ensuring that hydrogen peroxide does not have time to harm our cells.

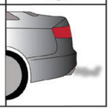
Enzymes are important in digestion, muscle function and many other essential aspects of human biology. Enzymes also have many industrial applications, such as fermentation to make wine and beer, 'biological' detergents, cheese manufacturing, contact lens cleaners, meat tenderisers, baby foods and paper manufacture.

• Suggested Films

- Rates of Reaction: Basic
- Collision Theory
- Transition Metals
- Fermentation

DIAGRAM 04:

Energy Changes
CHEMISTRY • REACTIONS • ENERGY CHANGES

Reaction	Purpose of reaction	Catalyst used	
The Haber process nitrogen + hydrogen ⇌ ammonia $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$	Making ammonia	iron	
The Contact process sulphur dioxide + oxygen ⇌ sulphur trioxide $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$	Making sulphuric acid	vanadium pentoxide	
Hydrogenation of vegetable oils $RCH = CHR + H_2 \rightarrow RCH_2CH_2R$	Making margarine	nickel	
The Ostwald process ammonia + oxygen ⇌ nitrogen monoxide + water $4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(l)$	Making nitric acid	platinum/rhodium	
Catalytic convertors nitric oxide + carbon monoxide → nitrogen + carbon dioxide $2CO(g) + 2NO(g) \rightarrow N_2(g) + 2CO_2(g)$	Reduces harmful car emissions	platinum/palladium/ rhodium	

Section 3: Redox Reactions and Electrolysis

• What are redox reactions?

A redox reaction is one in which one substance is reduced and another substance is oxidised. One meaning of reduction is **removal of oxygen**, and oxidation is **reaction with oxygen**. Redox reactions are often used to make metals from their ores, for example when tin ore (tin(IV) oxide or cassiterite) is heated with carbon:

tin(IV) oxide + carbon → tin + carbon dioxide

$$SnO_2(s) + C(s) \rightarrow Sn(s) + CO_2(g)$$

$$\Delta H = +187.2 \text{ kJmol}^{-1}$$

In this case, tin(IV) oxide has lost oxygen, so it is reduced. Carbon has gained oxygen, so it is oxidised.

Substances that can remove oxygen are called **reducing agents**. Carbon, in the form of charcoal, was one of the first reducing agents to be discovered, and was used during the Bronze Age to make tin and copper. In the blast furnace, carbon monoxide is the reducing agent. It removes oxygen from iron(III) oxide and is oxidised to carbon dioxide:

• Suggested Films
- Redox Reactions
- Rates of Reaction: Basic
- Collision Theory
- Transition Metals
- The Elements: Iron
- Oxidation
- The Elements: Copper
- The Elements: Iron
- The Elements: Potassium
- Reactivity Series

• Worksheet Question
- Question 4

iron(III) oxide + carbon monoxide → iron + carbon dioxide

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$$

$$\Delta H = -24.8 \text{ kJmol}^{-1}$$

Redox reactions may be endothermic or exothermic, but due to the high activation energies of the reactions, the reactants always need to be heated to a high temperature before they can begin. Many of the metals we use, including tin, iron, zinc, lead and copper can be extracted from their ores. Whereas the more reactive metals, such as sodium, potassium, magnesium and aluminium cannot be extracted in this way, and instead are extracted using a technique called electrolysis.

• What is electrolysis?

Electrolysis is the process whereby **electricity is used to decompose a compound.** The compound that is decomposed is called the electrolyte. Electrolytes are made of **ions**, and are usually compounds of a metal and a non-metal, such as lead bromide $PbBr_2$. Electrolytes can only conduct electricity when they are either molten or in solution, as the current is carried by the moving ions. Ions can only move through liquids. In solid lead(II) bromide the lead(II) ions Pb^{2+} and bromide ions Br^- are fixed in position and cannot move, therefore no current can flow through the solid. In molten lead(II) bromide the ions are able to move freely, and it is decomposed into liquid lead and bromine gas:

$$PbBr_2(l) \rightarrow Pb(l) + Br_2(g)$$

• Suggested Films

- Electrolysis
- Oxidation Reactions
- Ionic Bonding
- Alkali Metals
- The Elements: Potassium
- The Elements: Lead
- The Halogens

• Worksheet Question

- Question 5

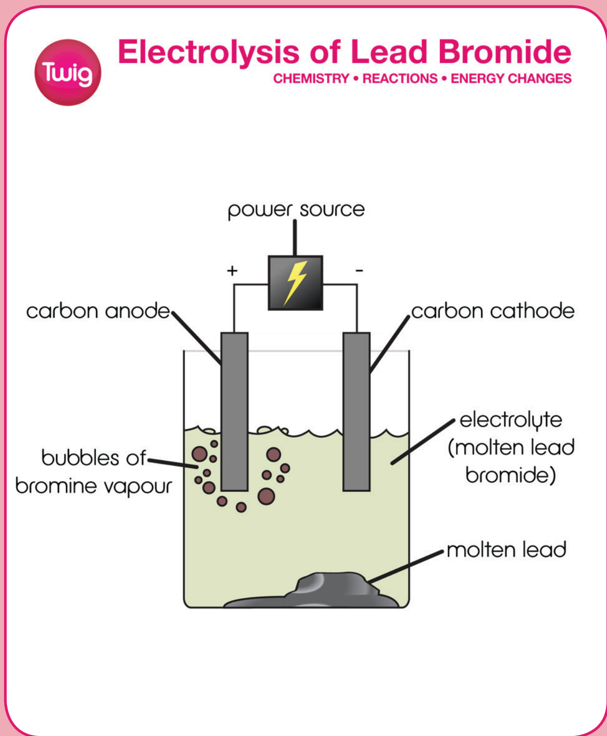
In order to carry out electrolysis, the electrolyte is put in a container and two rods called electrodes are inserted into the electrolyte. The positive electrode is called the anode; the negative electrode is called the cathode. The electrodes may be made of a metal or carbon, as they have to be able to conduct electricity. If the electrolyte is molten, the electrodes need to have a fairly high melting point. The anode and cathode are connected to a power source such as a cell (battery). Provided the electrolyte is a liquid (molten or in solution), the ions will start to drift towards the cathode or anode, and chemical reactions will take place. Positive ions move towards the cathode; negative ions move towards the anode.

Electrolysis is an endothermic process, as energy has to be put in to decompose the compound. During the electrolysis of molten lead(II) bromide, the lead(II) ions drift towards the cathode, where they gain electrons and become molten lead metal:

(I) REDUCTION occurs at the cathode: $Pb^{2+} + 2e^- \rightarrow Pb$

The bromide ions drift towards the anode, where they lose electrons, becoming bromine atoms, which immediately pair up to form diatomic bromine molecules:

DIAGRAM 05:



(II) OXIDATION occurs at the anode: $2Br^- \rightarrow Br_2 + 2e^-$

The above equations are called half-equations because they only show **either** the oxidation process **or** the reduction process. Half-equations always include electrons on one side or the other of the arrow. In electrolysis reactions it is helpful to use a different definition of oxidation and reduction from the one in the previous section. **We say oxidation involves losing electrons and reduction involves gaining electrons.** This is sometimes given the acronym **OIL- RIG (Oxidation Is Loss - Reduction Is Gain)**. In this case, lead(II) ions gain electrons, so they are reduced, and bromide ions lose electrons, so they are oxidised. By combining equations (I) and (II) we get the overall redox reaction during electrolysis:

$$Pb^{2+} + 2Br^- \rightarrow Pb + Br_2$$

• How is aluminium extracted?

Before 1800, reactive metals such as potassium and sodium had never been extracted. Traditional reducing agents such as carbon and carbon monoxide were simply unable to pull oxygen away from their oxides. Only with the introduction of powerful electrical batteries in the early 19th century were these very reactive metals extracted. Humphry Davy (1778–1829) was the first to use electrolysis to isolate the metals sodium, potassium, calcium, magnesium and barium. Aluminium is the most abundant metal in the Earth's crust. It is also highly reactive, which means that electrolysis is needed to extract it from its ore. Aluminium occurs as the mineral bauxite, which is impure aluminium oxide. The first part of the extraction process involves removing the impurities to leave the pure white powder, aluminium oxide Al_2O_3 or alumina. Aluminium oxide has very high melting point of 2054°C, and so it requires a large amount of energy to melt. In the early 19th century, the high energy costs of the process made aluminium almost as costly as silver. The breakthrough came when Charles Hall (1863–1914) found that the mineral cryolite Na_3AlF_6 melted at around 1000°C and was a good solvent for aluminium oxide. The French scientist, Paul Héroult (1863–1914), made a similar discovery at almost the same time. The Hall-Héroult process is still the basis for aluminium extraction today. Their discovery meant that aluminium oxide could be dissolved in molten cryolite, so the aluminium ions and oxide ions are able to move freely, and electrolysis can take place. In the modern process, the molten cryolite with dissolved aluminium oxide, is held in a steel vessel at about 1000°C; the cathode and anodes are made of carbon.

At the cathode, aluminium ions gain electrons (they are reduced) and become aluminium atoms, forming liquid aluminium which is poured off and allowed to cool:

$$Al^{3+} + 3e^- \rightarrow Al \qquad \text{REDUCTION}$$

At the anode, oxide ions lose electrons (they are oxidised) making oxygen gas:

$$2O^{2-} \rightarrow O_2 + 4e^- \quad \text{OXIDATION}$$

The oxygen gas reacts with the carbon anode, making carbon dioxide:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

The anodes burn away and have to be replaced at regular intervals.

The overall reaction is highly endothermic, because the ionic bonding in aluminium oxide is extremely strong, and large amounts of energy are needed to break up the ionic structure and extract aluminium metal.

$$2Al_2O_3 \rightarrow 4Al(s) + 3O_2(g)$$
$$\Delta H = +3315.4 \text{ kJmol}^{-1}$$

• Suggested Films

- Electrolysis
- Ionic Bonding
- Alkali Metals
- The Elements: Potassium
- Reactivity Series
- Extraction of Aluminium

• Worksheet Questions

- Questions 6 and 7

• Worksheet

Q1. Which of these processes involves chemical reactions that are exothermic?

(a) Propane gas burning, **(b)** Cooking an egg, **(c)** A firework exploding

(a) ____________________

(b) ____________________

(c) ____________________

Q2. What is the enthalpy change for the reaction $H_2(g) + Br_2(g) \rightarrow 2HBr(g)$?

H-H(g) → H(g) + H(g) **ΔH = + 436 kJmol^{-1}**
Br-Br(g) → Br(g) + Br(g) **ΔH = +193 kJmol^{-1}**
H-Br(g) → H(g) + Br(g) **ΔH = + 366 kJmol^{-1}**

Use the data above to calculate the enthalpy change for the reaction, using the sign and unit. Remember: ΔH = (sum of enthalpy changes when bonds are broken) – (sum of enthalpy changes when bonds are made)

Q3. Why does the reaction of carbon with oxygen have a high activation energy, even though the reaction is highly exothermic?

• Worksheet

Q4. In the reaction $2CuO(s) + C(s) \rightarrow 2Cu(s) + CO_2(g)$ what is oxidised, what is reduced, and what is the reducing agent?

Q5. Complete the half-equations below, to show the reactions which take place during the electrolysis of magnesium chloride. Then answer what is oxidised, what is reduced, and what is the overall reaction?

At the cathode: Mg^{2+} _______$\rightarrow$ Mg At the anode: ___Cl^- $\rightarrow$ Cl_2 ___

Q6. Before the electrolysis of aluminium oxide became feasible, a possible method for making aluminium would have been: $6Na(s) + Al_2O_3(s) \rightarrow 2Al(s) + 3Na_2O(s)$
Why would this have been an expensive method?

Q7. Aluminium is highly reactive. Why is it possible to use aluminium foil to wrap food and to make saucepans?

• Worksheet Answers

Q1. Which of these processes involves chemical reactions that are exothermic?

(a) Propane gas burning, **(b)** Cooking an egg, **(c)** A firework exploding

(a) involves exothermic combustion reactions

(b) involves heat being taken in, therefore it is an endothermic process

(c) involves exothermic combustion reactions

Q2. What is the enthalpy change for the reaction $H_2(g) + Br_2(g) \rightarrow 2HBr(g)$?

H-H(g) → H(g) + H(g) **ΔH = + 436 kJmol^{-1}**
Br-Br(g) → Br(g) + Br(g) **ΔH = +193 kJmol^{-1}**
H-Br(g) → H(g) + Br(g) **ΔH = + 366 kJmol^{-1}**

Use the data above to calculate the enthalpy change for the reaction, using the sign and unit. Remember: ΔH = (sum of enthalpy changes when bonds are broken) – (sum of enthalpy changes when bonds are made)

436 + 193 – (2 x 366) = -103 kJmol^{-1}

Q3. Why does the reaction of carbon with oxygen have a high activation energy, even though the reaction is highly exothermic?

Carbon (charcoal) has a giant structure of atoms in which there are very strong (C-C) covalent bonds between the carbon atoms, and a lot of energy is needed to break the carbon atoms apart. Oxygen O_2 molecules have strong covalent (O=O) bonds which also have to be broken before the reaction can begin. The result is that the activation energy for the reaction is very high, so the reaction with oxygen is very slow unless the solid carbon is first heated strongly.

• Worksheet Answers

Q4. In the reaction $2CuO(s) + C(s) \rightarrow 2Cu(s) + CO_2(g)$ what is oxidised, what is reduced, and what is the reducing agent?

C is oxidised to CO_2
CuO is reduced to Cu
Carbon is the reducing agent.

Q5. Complete the half-equations below, to show the reactions which take place during the electrolysis of magnesium chloride. Then answer what is oxidised, what is reduced, and what is the overall reaction?

At the cathode: $Mg^{2+} + 2e^- \rightarrow Mg$ At the anode: $2Cl^- \rightarrow Cl_2 + 2e^-$

Magnesium ions are reduced to magnesium atoms at the cathode.
Chloride ions are oxidised to chlorine molecules at the anode.
The overall equation is $Mg^{2+} + 2Cl^- \rightarrow Mg + Cl_2$

**Q6. Before the electrolysis of aluminium oxide became feasible, a possible method for making aluminium would have been: $6Na(s) + Al_2O_3(s) \rightarrow 2Al(s) + 3Na_2O(s)$
Why would this have been an expensive method?**

Sodium itself can only be obtained by electrolysis, so it would have been an expensive metal. The reaction would also have needed a high temperature due to the very high melting point of aluminium oxide.

Q7. Aluminium is highly reactive. Why is it possible to use aluminium foil to wrap food and to make saucepans?

Aluminium has a very thin, coherent layer of aluminium oxide on its surface, which acts as a barrier, and prevents air and water penetrating to the aluminium metal beneath. It prevents aluminium reacting as violently as its place in the Reactivity Series might suggest.

• Quizzes

Oxygen and Combustion

Basic

• Which of these is NOT essential for combustion?

A – initial source of heat
B – fuel
C – source of oxygen
D – water

• When a substance is burned, the products always include compounds of

A – nitrogen
B – oxygen
C – carbon
D – sulphur

• Lavoisier discovered that when tin was heated it

A – gained mass
B – lost mass
C – changed colour
D – evaporated

• The correct formula for oxygen molecules is

A – O
B – O_2
C – O_3
D – O_4

Advanced

• When iron wool is heated, it gains mass because it

A – combines with oxygen
B – rusts
C – gives out light
D – expands

• When iron burns in air it makes

A – water
B – iron oxide
C – rust
D – iron nitride

• Which of these compounds could NOT be a product of combustion?

A – CO
B – CO_2
C – CH_4
D – SO_2

• When hydrogen burns in oxygen it makes

A – hydrogen peroxide
B – carbon dioxide
C – water
D – carbon monoxide

Electrolysis

Basic

• When Humphry Davy electrolysed potassium hydroxide, the colour of the flame he noticed was

A – red

B – lilac

C – blue

D – green

• The negative electrode is called

A – cathode

B – anode

C – ion

D – electrolysis

• Electrolysis of potassium hydroxide was not possible before 1800 because

A – potassium hydroxide had not been discovered

B – electrolysis was too expensive

C – electrolysis was too dangerous

D – electric batteries were not available

• A metal that is still extracted using electrolysis is

A – iron

B – tungsten

C – aluminium

D – gold

Advanced

• During electrolysis of potassium hydroxide, the current in the molten compound is carried by moving

A – ions

B – atoms

C – electrons

D – protons

• Which of these compounds is NOT made of ions?

A – potassium hydroxide KOH

B – sodium hydroxide NaOH

C – methane CH_4

D – magnesium oxide MgO

• During electrolysis of potassium hydroxide, which of these products is NOT made?

A – oxygen at the cathode

B – water at the anode

C – potassium at the cathode

D – oxygen at the anode

• Which of these statements about electrolysis is NOT true?

A – it is a chemical reaction

B – one or more elements are made during the process

C – energy is taken in

D – it can be used to make noble gases

Oxidation Reactions

Basic

- Which of these is NOT an oxidation reaction?

  A – rusting

  B – combustion

  C – digestion

  D – respiration

- The surface of aluminium has a thin layer of

  A – aluminium nitride

  B – aluminium oxide

  C – aluminium sulphide

  D – aluminium carbonate

- The chemical name for rust is

  A – iron (II) oxide

  B – iron (III) oxide

  C – hydrated iron (II) oxide

  D – hydrated iron (III) oxide

- Aluminium is useful for wrapping food because it

  A – is reactive

  B – is unreactive

  C – has a low density

  D – conducts electricity

Advanced

- Rusting occurs when iron is exposed to

  A – water

  B – water and oxygen

  C – oxygen

  D – water and nitrogen

- The surface layer of aluminium oxide means that aluminium

  A – is rather unreactive with water and oxygen

  B – does not conduct heat

  C – is brittle

  D – does not conduct electricity

- A metal that does NOT react with oxygen at all is

  A – magnesium

  B – sodium

  C – gold

  D – copper

- Which of these statements about oxidation of metals is NOT true?

  A – the metal gains mass when it is oxidised

  B – metal oxides are formed during oxidation

  C – metal oxides are ionic solids

  D – oxidation of metals is endothermic

• Answers

Oxygen and Combustion

Basic

• Which of these is NOT essential for combustion?

A – initial source of heat

B – fuel

C – source of oxygen

$\boxed{\text{D – water}}$

• When a substance is burned, the products always include compounds of

A – nitrogen

$\boxed{\text{B – oxygen}}$

C – carbon

D – sulphur

• Lavoisier discovered that when tin was heated it

$\boxed{\text{A – gained mass}}$

B – lost mass

C – changed colour

D – evaporated

• The correct formula for oxygen molecules is

A – O

$\boxed{\text{B – } O_2}$

C – O_3

D – O_4

Advanced

• When iron wool is heated, it gains mass because it

$\boxed{\text{A – combines with oxygen}}$

B – rusts

C – gives out light

D – expands

• When iron burns in air it makes

A – water

$\boxed{\text{B – iron oxide}}$

C – rust

D – iron nitride

• Which of these compounds could NOT be a product of combustion?

A – CO

B – CO_2

$\boxed{\text{C – } CH_4}$

D – SO_2

• When hydrogen burns in oxygen it makes

A – hydrogen peroxide

B – carbon dioxide

$\boxed{\text{C – water}}$

D – carbon monoxide

Electrolysis

Basic

• When Humphry Davy electrolysed potassium hydroxide, the colour of the flame he noticed was

A – red

B – lilac

C – blue

D – green

• The negative electrode is called

A – cathode

B – anode

C – ion

D – electrolysis

• Electrolysis of potassium hydroxide was not possible before 1800 because

A – potassium hydroxide had not been discovered

B – electrolysis was too expensive

C – electrolysis was too dangerous

D – electric batteries were not available

• A metal that is still extracted using electrolysis is

A – iron

B – tungsten

C – aluminium

D – gold

Advanced

• During electrolysis of potassium hydroxide, the current in the molten compound is carried by moving

A – ions

B – atoms

C – electrons

D – protons

• Which of these compounds is NOT made of ions?

A – potassium hydroxide KOH

B – sodium hydroxide NaOH

C – methane CH_4

D – magnesium oxide MgO

• During electrolysis of potassium hydroxide, which of these products is NOT made?

A – oxygen at the cathode

B – water at the anode

C – potassium at the cathode

D – oxygen at the anode

• Which of these statements about electrolysis is NOT true?

A – it is a chemical reaction

B – one or more elements are made during the process

C – energy is taken in

D – it can be used to make noble gases

Oxidation Reactions

Basic

• Which of these is NOT an oxidation reaction?

A – rusting

B – combustion

$\boxed{\text{C – digestion}}$

D – respiration

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