10 Metals

10.1.1 Properties of Metals

Physical & Chemical Properties of Metals

PROPERTY	REASON
HIGH MELTING AND BOILING POINT	THERE ARE MANY STRONG METALLIC BONDS IN GIANT METALLIC STRUCTURES SO LARGE AMOUNTS OF HEAT ENERGY ARE NEEDED TO OVERCOME FORCES AND BREAK THESE BONDS.
GOOD CONDUCTORS OF ELECTRICITY AND HEAT	METALS ARE GOOD CONDUCTORS BECAUSE OF THE FREE ELECTRONS THAT ARE AVAILABLE TO MOVE AND CARRY CHARGE. WHEN A METAL IS USED IN AN ELECTRICAL CIRCUIT, ELECTRONS ENTERING ONE END OF THE METAL CAUSE A DELOCALISED ELECTRON TO DISPLACE ITSELF FROM THE OTHER END. HENCE ELECTRONS CAN FLOW SO ELECTRICITY IS CONDUCTED.
MALLEABLE AND DUCTILE	LAYERS OF POSITIVE IONS CAN EASILY SLIDE OVER ONE ANOTHER AND TAKE UP DIFFERENT POSITIONS. THIS DOES NOT DISRUPT THE METALLIC BONDING AS THE VALENCE ELECTRONS DO NOT BELONG TO ANY PARTICULAR METAL ATOM AND SO THEY CAN MOVE WITH THE LAYERS OF POSITIVE IONS, MAINTAINING THE ELECTROSTATIC FORCES. THE METALLIC BONDS ARE THUS NOT BROKEN AND AS A RESULT METALLIC BONDS ARE STRONG BUT FLEXIBLE. THEREFORE, THEY CAN BE HAMMERED INTO DIFFERENT SHAPES WITHOUT BREAKING.

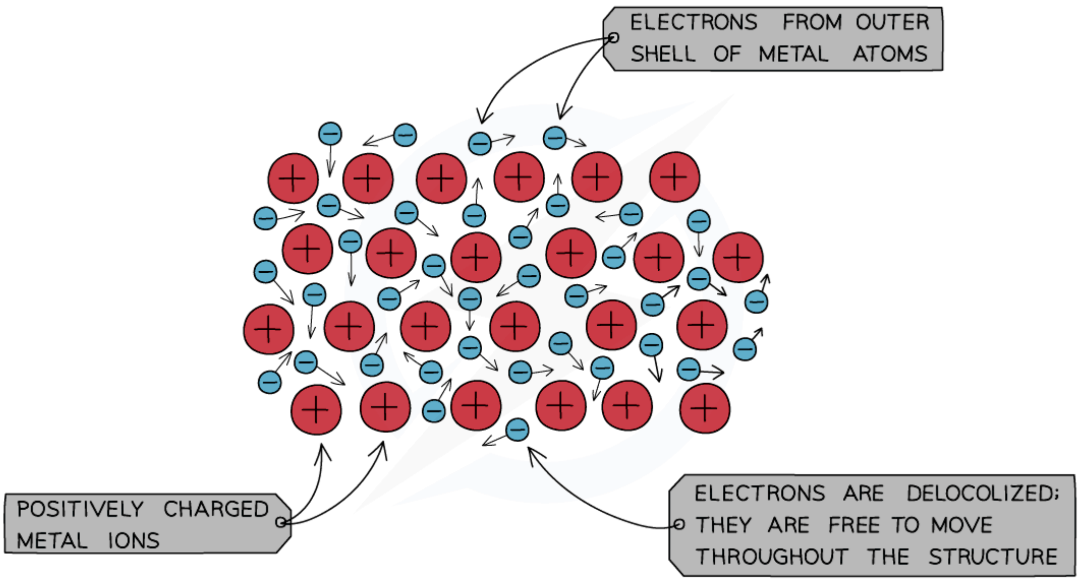


Diagram showing bonding and structure in metals

General chemical properties of metals

- The chemistry of metals is studied by analysing their reactions with water, dilute acid and oxygen
- Based on these reactions, a reactivity series of metals can be produced

Reactivity with water

- Some metals react with water, either warm or cold, or with steam
- Metals that react with cold water form a metal hydroxide and hydrogen gas, for example calcium:

$$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$$

- Metals that react with steam form metal oxide and hydrogen gas, for example zinc:

$$Zn + H_2O \rightarrow ZnO + H_2$$

Reactivity with acids

- Most metals react with dilute acids such as HCl
- When acids and metals react, the hydrogen atom in the acid is replaced by the metal atom to produce a salt and hydrogen gas, for example iron:

$$Fe + 2HCI \rightarrow FeCl_2 + H_2$$

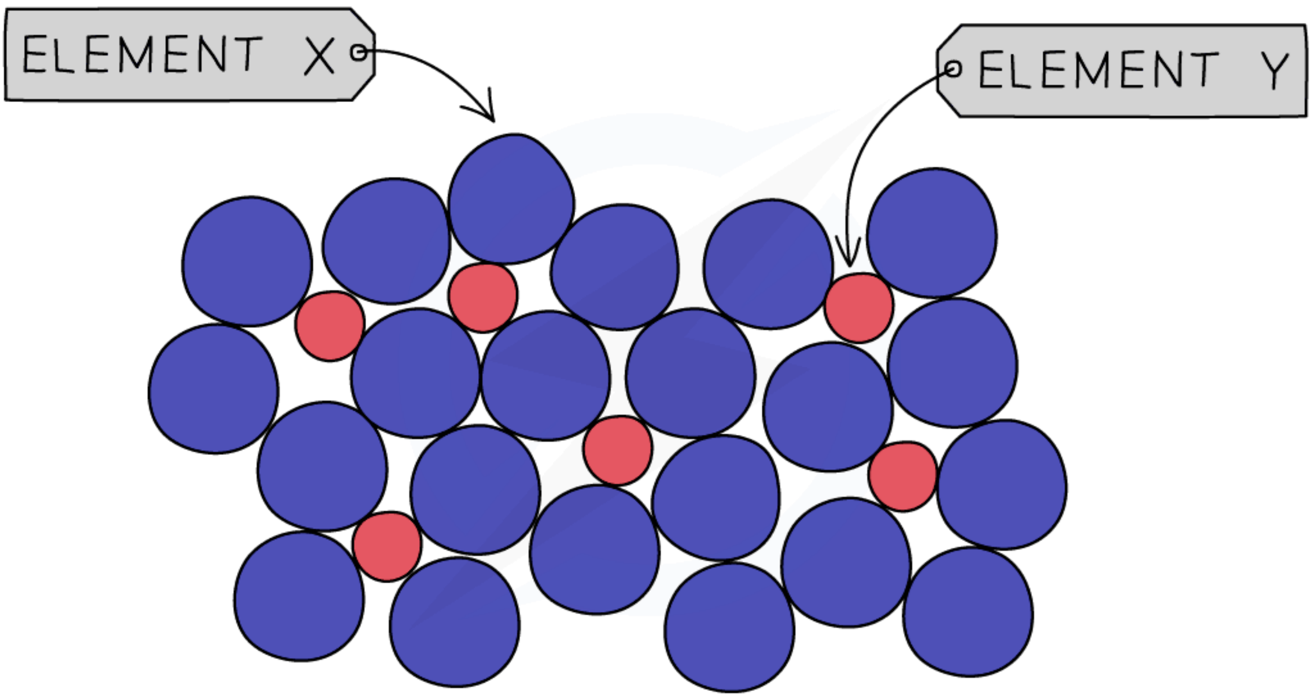
Reactivity with oxygen

- Unreactive metals such as gold and copper do not react with acids
- Some reactive metals such as the alkali metals react with oxygen
- Copper and iron can also react with oxygen although much more slowly
- When metals react with oxygen a metal oxide is formed, for example copper:

$$2Cu + O_2 \rightarrow 2CuO$$

Structure & Uses of Alloys

- An alloy is a **mixture** of two or more metals or a metal and a nonmetal
- Alloys often have **properties** that can be very **different** to the metals they contain, for example they can have more **strength**, **hardness** or **resistance** to **corrosion** or extreme **temperatures**
- Alloys contain atoms of **different** sizes, which **distorts** the normally regular arrangements of atoms in metals
- This makes it more difficult for the layers to **slide** over each other, so alloys are usually much harder than the pure metal



The regular arrangement of a metal lattice structure is distorted in alloys

Common alloys and their uses

- Brass is an alloy of copper and zinc and is much **stronger** than either metal
- Alloys of iron with tungsten are extremely **hard** and resistant to high **temperatures**
- Alloys of iron mixed with chromium or nickel are resistant to **corrosion**
- Aluminium is mixed with copper, manganese and silicon for aircraft body production as the alloy is **stronger** but still has a low **density**

Exam Tip

Alloys are mixtures of substances, they are not chemically combined and an alloy is not a compound.

10.1.2 Reactivity Series

The Reactivity Series

- The chemistry of the metals is studied by analysing their reactions with water, dilute acid and oxygen
- Based on these reactions a reactivity series of metals can be produced
- The series can be used to place a group of metals in **order of reactivity** based on the observations of their reactions with water, acid and oxygen

METAL	REACTION WITH WATER	REACTION WITH ACID	REACTION WITH OXYGEN
MOST REACTIVE			
POTASSIUM	REACTS VIOLENTLY	REACTS VIOLENTLY	REACTS QUICKLY IN AIR
SODIUM	REACTS QUICKLY	REACTS VIOLENTLY	REACTS QUICKLY IN AIR
CALCIUM	REACTS LESS STRONGLY	REACTS VIGOROUSLY	REACTS READILY
MAGNESIUM		REACTS VIGOROUSLY	REACTS READILY
ZINC		REACTS LESS STRONGLY	REACTS
IRON		REACTS LESS STRONGLY	REACTS
HYDROGEN			REACTS
COPPER			REACTS
LEAST REACTIVE			

Carbon and the reactivity series mnemonic

- Carbon is an important element and has its own place on the reactivity series
- Its use in the extraction of metals from their oxides is discussed in this section but a more complete reactivity series with an accompanying mnemonic to help you memorise it is below

The reactivity series mnemonic

- "Please send lions, cats, monkeys and cute zebras into hot countries signed Gordon"

METAL	ABBREVIATION
MOST REACTIVE	
POTASSIUM	P – PLEASE
SODIUM	S – SEND
LITHIUM	L – LIONS,
CALCIUM	C – CATS,
MAGNESIUM	M – MONKEYS,
ALUMINIUM	A – AND
CARBON	C – CUTE
ZINC	Z – ZEBRAS
IRON	I – INTO
HYDROGEN	H – HOT
COPPER	C – COUNTRIES
SILVER	S – SIGNED
GOLD	G – GORDON
LEAST REACTIVE	

Extended Only

Reactions with Aqueous Ions & Oxides

- The reactivity of metals increases going up the reactivity series
- This means that a more reactive metal can displace a less reactive metal from its oxide by heating

Example: Copper(II) Oxide

- It is possible to reduce copper(II) oxide by heating it with magnesium
- As magnesium is above copper in the reactivity series, magnesium is more reactive so can displace copper
- The reducing agent in the reaction is magnesium:

$$CuO\ (s) + Mg\ (s) \rightarrow Cu\ (s) + MgO\ (s)$$

Other common reactions

MIXTURE	PRODUCTS	EQUATION FOR REACTION
IRON (III) OXIDE AND ALUMINIUM (THERMITE REACTION)	IRON AND ALUMINIUM OXIDE	$Fe_2O_3 + 2Al \rightarrow 2Fe + Al_2O_3$
SODIUM OXIDE AND MAGNESIUM	NO REACTION AS SODIUM IS ABOVE MAGNESIUM	–
SILVER OXIDE AND COPPER	SILVER AND COPPER(II) OXIDE	$Ag_2O + Cu \rightarrow 2Ag + CuO$
ZINC OXIDE AND CALCIUM	ZINC AND CALCIUM OXIDE	$ZnO + Ca \rightarrow Zn + CaO$
LEAD (II) OXIDE AND SILVER	NO REACTION AS LEAD IS MORE REACTIVE THAN SILVER	–

Displacement reactions between metals and aqueous solutions of metal salts

- Any metal will displace another metal that is **below** it in the reactivity series from a solution of one of its salts
- This is because more reactive metals lose electrons and form ions more readily than less reactive metals, making them better **reducing agents**
- The less reactive metal is a better electron acceptor than the more reactive metal, thus the less reactive metal is reduced. (OIL-RIG: reduction is gain of electrons)

Example: Zinc and copper(II) sulfate

- As Zinc is above copper in the reactivity series, zinc is more reactive so can displace copper from copper(II) sulfate solution:

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

Other Common Reactions

MIXTURE	PRODUCTS	EQUATION FOR REACTION
MAGNESIUM AND IRON(II) SULFATE	MAGNESIUM SULFATE AND IRON	$Mg + FeSO_4 \longrightarrow MgSO_4 + Fe$
ZINC AND SODIUM CHLORIDE	NO REACTION AS SODIUM IS ABOVE ZINC	–
LEAD AND SILVER NITRATE	LEAD (II) NITRATE AND SILVER	$Pb + 2AgNO_3 \longrightarrow Pb(NO_3)_2 + 2Ag$
COPPER AND CALCIUM CHLORIDE	NO REACTION AS CALCIUM IS MORE REACTIVE THAN COPPER	–
IRON AND COPPER(II) SULFATE	IRON (II) SULFATE AND COPPER	$Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$

Extended Only

Heating Metal Hydroxides, Carbonates & Nitrates

Thermal decomposition reactions

- Some compounds **decompose** or **breakdown** when they are heated to sufficiently high temperatures
- These reactions are called thermal decomposition reactions
- A common example is the thermal decomposition of calcium carbonate (limestone), which occurs at temperatures above 800ºC:

$$CaCO_3 \rightarrow CaO + CO_2$$

Thermal decomposition of metal hydroxides

- Most metal hydroxides undergo thermal decomposition
- **Water** and the corresponding **metal oxide** are the products formed, for example zinc hydroxide thermally decomposes as follows:

$$Zn(OH)_2 \rightarrow ZnO + H_2O$$

- Group II metal hydroxides decompose similarly but the Group I hydroxides (apart from lithium) do **not** decompose due to their having a higher thermal stability

Thermal decomposition of metal carbonates

- Most of the metal carbonates and hydrogen carbonates undergo thermal decomposition
- The **metal oxide** and **carbon dioxide** are the products formed, for example magnesium carbonate thermally decomposes as follows:

$$MgCO_3 \rightarrow MgO + CO_2$$

- Group I carbonates (again apart from lithium carbonate) do **not** decompose when heated
- This is due to the high thermal stability of reactive metals; the more reactive the metal then the more difficult it is to decompose its carbonate
- $CuCO_3$ for example is relatively easy to thermally decompose but K_2CO_3 does not decompose

Thermal decomposition of metal nitrates

- All of the metal nitrates decompose when they are heated
- Group I nitrates decompose forming the **metal nitrite** and **oxygen**, for example sodium nitrate decomposes as follows:

$$2NaNO_3 \rightarrow 2NaNO_2 + O_2$$

- Most other metal nitrates form the corresponding **metal oxide**, **nitrogen dioxide** and **oxygen** when heated, for example copper nitrate:

$$2Cu(NO_3)_2 \rightarrow 2CuO + 4NO_2 + O_2$$

Aluminium and its apparent lack of reactivity

- Aluminium is a curious metal in terms of its reactivity
- It is placed **high** on the reactivity series but it doesn´t react with water or acids
- This is because the surface of aluminium metal reacts with oxygen in the air forming a protective coating of **aluminium oxide**:

$$4Al + 3O_2 \rightarrow 2Al_2O_3$$

- The aluminium oxide layer is **tough**, **unreactive** and **resistant** to **corrosion**
- It adheres very strongly to the aluminium surface and protects it from reaction with other substances, hence making it appear unreactive

Exam Tip

For the thermal decomposition reactions, you will need to be able to describe how the Group I nitrates differ from the other metals.

You should be able to write out the balanced symbol equations for these reactions.

10.2.1 Extraction of Metals

Obtaining Metals from their Ores

Extraction of ores from the Earth's crust

- The Earth's crust contains metals and metal compounds such as gold, iron oxide and aluminium oxide
- When found in the Earth, these are often mixed with other substances
- To be useful, the metals have to be extracted from their ores through processes such as electrolysis, using a blast furnace or by reacting with more reactive material
- The extraction of metals is a reduction process
- Unreactive metals do not have to be extracted as they are often found as the uncombined element as they do not easily react with other substances

Extraction of metal and the reactivity series

- The position of the metal on the reactivity series influences the method of extraction
- Those metals placed higher up on the series (above carbon) have to be extracted using electrolysis
- Metals lower down on the series can be extracted by heating with carbon

The reactivity series and extraction of metals

<table>
<tr><th>METAL</th><th>ABBREVIATION</th></tr>
<tr><td colspan="2">MOST REACTIVE</td></tr>
<tr><td>POTASSIUM</td><td rowspan="6">EXTRACTED BY ELECTROLYSIS OF THE MOLTEN CHLORIDE OR MOLTEN OXIDE
LARGE AMOUNTS OF ELECTRICITY REQUIRED SO EXPENSIVE PROCESS</td></tr>
<tr><td>SODIUM</td></tr>
<tr><td>LITHIUM</td></tr>
<tr><td>CALCIUM</td></tr>
<tr><td>MAGNESIUM</td></tr>
<tr><td>ALUMINIUM</td></tr>
<tr><td>CARBON</td><td></td></tr>
<tr><td>ZINC</td><td rowspan="4">EXTRACTED BY HEATING WITH A REDUCING AGENT SUCH AS CARBON OR CARBON MONOXIDE IN A BLAST FURNACE
CHEAP PROCESS AS CARBON IS CHEAP AND CAN BE SOURCE OF HEAT AS WELL</td></tr>
<tr><td>IRON</td></tr>
<tr><td>HYDROGEN</td></tr>
<tr><td>COPPER</td></tr>
<tr><td>SILVER</td><td rowspan="2">FOUND AS PURE ELEMENTS</td></tr>
<tr><td>GOLD</td></tr>
<tr><td colspan="2">LEAST REACTIVE</td></tr>
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Extraction of Iron from Hematite

The extraction of iron in the blast furnace

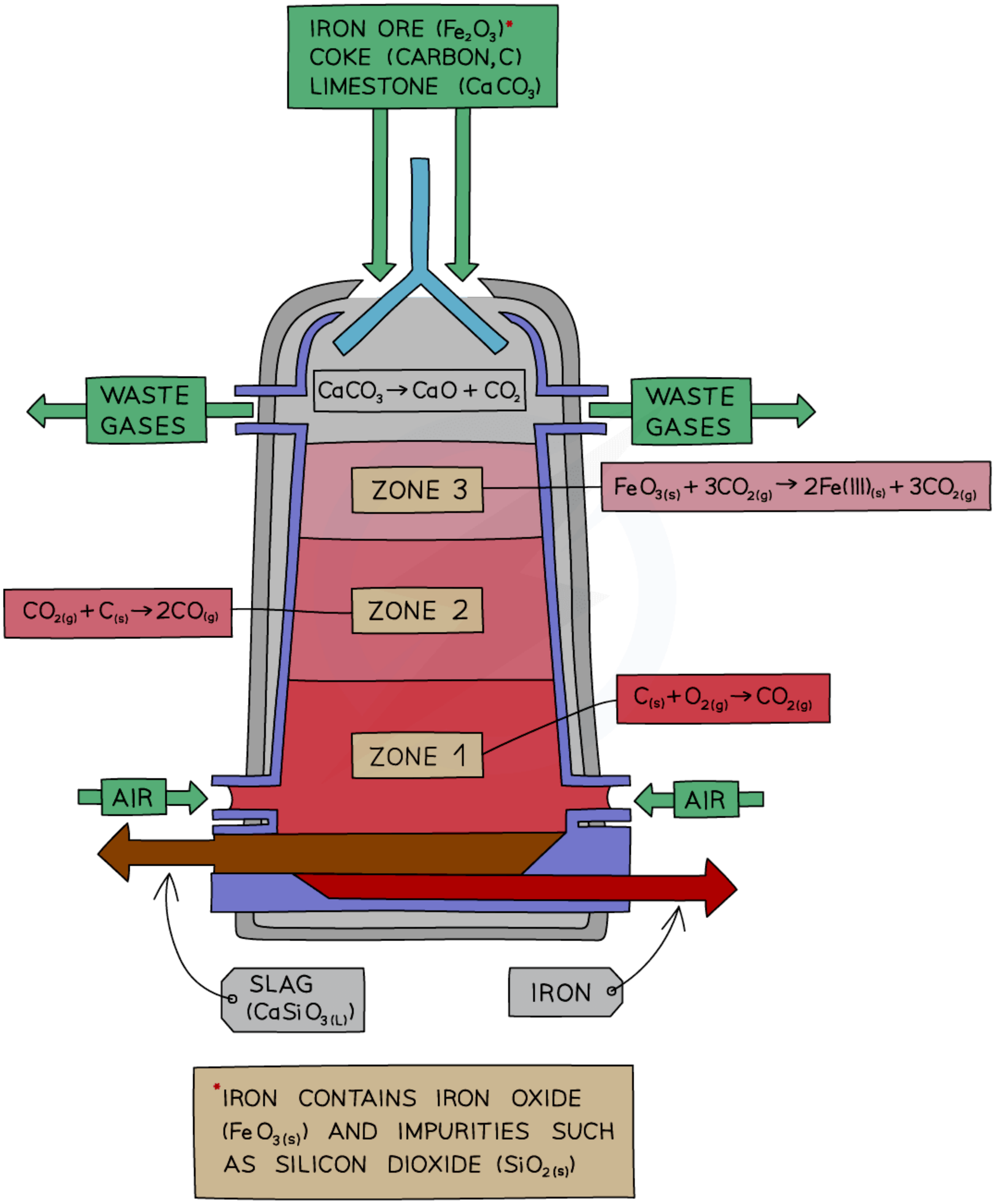

Diagram showing the carbon extraction of iron

Raw Materials: Iron Ore (Haematite), Coke, Limestone and Air

Explanation:

- Iron Ore, Coke and Limestone are mixed together and fed into the top of the blast furnace. Hot air is blasted into the bottom of the blast furnace
- **Zone 1**
  - Coke is used as the starting material.
  - It is an impure carbon and it burns in the hot air blast to form carbon dioxide.
  - This is a strongly exothermic reaction:

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

- **Zone 2**
  - At the high temperatures in the furnace, carbon dioxide reacts with coke to form carbon monoxide:

$$CO_2\ (g) + C\ (s) \rightarrow 2CO\ (g)$$

- **Zone 3**
  - Carbon Monoxide (the reducing agent) reduces the Iron (III) Oxide in the Iron Ore to form Iron
  - This will melt and collect at the bottom of the furnace, where it is tapped off:

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

- Limestone is added to the furnace to remove impurities in the ore.
- The Calcium Carbonate in the limestone decomposes to form calcium Oxide:

$$CaCO_3\ (s) \rightarrow CaO\ (s) + CO_2\ (g)$$

- The Calcium Oxide reacts with the Silicon Dioxide, which is an impurity in the Iron Ore, to form Calcium Silicate
- This melts and collects as a molten slag floating on top of the molten Iron, which is tapped off separately:

$$CaO\ (s) + SiO_2\ (s) \rightarrow CaSiO_3\ (l)$$

The Conversion of Iron into Steel

Making steel from iron

- Molten iron is an alloy of 96% iron, with carbon, phosphorus, silicon and sulfur impurities
- It is too brittle for most uses, so most of it is converted into steel by removing some of the impurities
- Not all of the carbon is removed as steel contains some carbon, the percentage of which depends on the use of the steel
- The molten iron is transferred to a **tilting furnace** where the conversion to steel takes place
- **Oxygen** and **powdered calcium oxide** are added to the iron
- The oxygen oxidises the carbon, phosphorus, silicon and sulfur to their oxides which are all **acidic**
- CO_2 and SO_2 are gaseous so escape from the furnace
- The acidic silicon and phosphorus oxides react with the powdered calcium oxide and from a slag which is mainly calcium silicate:

$$SiO_2(l) + CaO(s) \rightarrow CaSiO_3(s)$$

- The slag floats on the surface of the molten iron and is removed.

Steel alloys

- The amount of carbon removed depends on the amount of oxygen used
- By carefully controlling the amount of carbon removed and subsequent addition of other metals such as **chromium**, **manganese** or **nickel**, the particular type of steel alloy is produced

TYPE OF STEEL	IRON ALLOYED WITH	USE	MOST IMPORTANT PROPERTY
STAINLESS STEEL	20% CHROMIUM AND 10% NICKEL	CUTLERY AND SINKS, CHEMICAL PLANTS	STRONG AND RESISTANT TO CORROSION
TUNGSTEN STEEL	5% TUNGSTEN	EDGES OF HIGH SPEED CUTTING TOOLS	TOUGH AND HARD AT VERY HIGH TEMPERATURES
MAGANESE STEEL	13% MAGANESE	DRILL BITS, SPRINGS	VERY TOUGH AND SPRINGLY

Aluminium Extraction & Benefits of Recycling

Extraction of aluminium

- Aluminium is a reactive metal which sits above carbon on the reactivity series.
- It cannot be extracted from its ore (bauxite) by carbon reduction, so electrolysis is used.

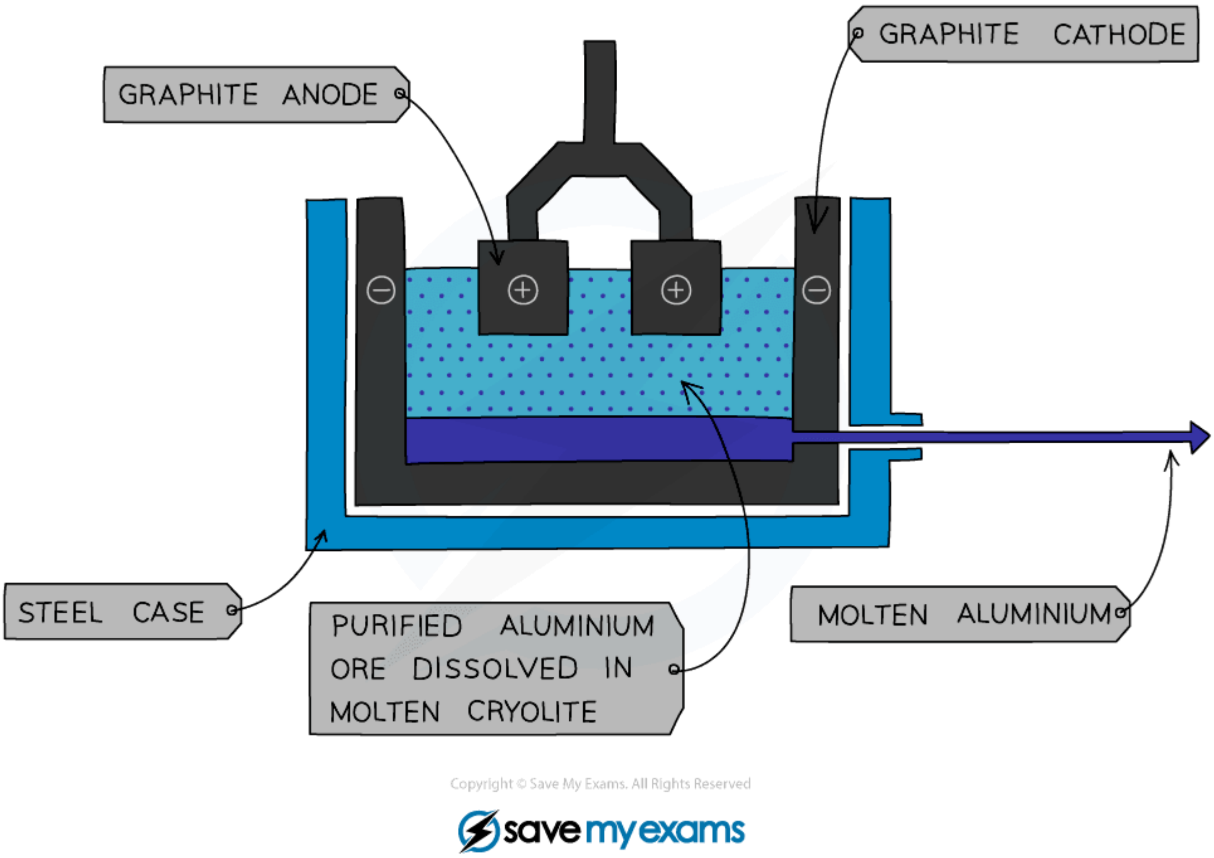

Diagram showing the extraction of aluminium by electrolysis

Recycling metals: iron, steel and aluminium

Advantages

- Raw materials are conserved (bauxite and haematite)
- Energy use is reduced, especially in the electrolysis of aluminium
- Less pollution is produced as both processes contribute to air pollution

Disadvantages

- More transport on roads carrying used metals to recycling centres
- Energy consumed in collecting materials and sorting them per material type

Extended Only

The Process of Aluminium Extraction by Electrolysis

Raw materials: Aluminium Ore (Bauxite)

Explanation:

- The Bauxite is first purified to produce Aluminium Oxide Al_2O_3
- Aluminium Oxide has a very high melting point so it is first dissolved in molten Cryolite producing an electrolyte with a lower melting point, as well as a better conductor of electricity than molten aluminium oxide. This also reduces expense considerably
- The electrolyte is a solution of aluminium oxide in molten cryolite at a temperature of about 1000 °C. The molten aluminium is siphoned off from time to time and fresh aluminium oxide is added to the cell. The cell operates at 5-6 volts and with a current of 100,000 amps. The heat generated by the huge current keeps the electrolyte molten
- A lot of electricity is required for this process of extraction, this is a major expense

Reaction at the negative electrode:

The Aluminium melts and collects at the bottom of the cell and is then tapped off:

$$Al^{3+} \quad + \quad 3e^- \quad \rightarrow \quad Al$$

Reaction at the positive electrode:

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

Some of the Oxygen produced at the positive electrode then reacts with the Graphite (Carbon) electrode to produce Carbon Dioxide Gas:

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

*This causes the carbon anodes to burn away, so they must be replaced regularly.

Extended Only

The Process of Zinc Extraction

- Zinc ore is called zinc blende, ZnS
- The zinc blende is first converted to zinc oxide by heating with air:

$$2ZnS + 3O_2 \rightarrow 2ZnO + 2SO_2$$

- The reducing agent is carbon monoxide which is formed inside the furnace through a series of reactions
- Carbon burns in a blast of very hot air to form carbon dioxide:

$$C + O_2 \rightarrow CO_2$$

- The carbon dioxide produced reacts with more coke to form carbon monoxide:

$$CO_2 + C \rightarrow 2CO$$

- The carbon monoxide is the reducing agent and reduces the zinc oxide to zinc:

$$ZnO(s) + CO(g) \rightarrow Zn(g) + CO_2(g)$$

- Note that the zinc produced is in the **gaseous** state
- This passes out of the furnace and is cooled and condensed in a tray placed at the top of the furnace
- This is a key difference between the extraction of iron and aluminium, both of which are collected at the **bottom** of the furnace / electrolytic cell in the **liquid** state

Uses of zinc

- Zinc is used in galvanising, the process of coating a metal such as iron or steel with a protective coating of zinc to prevent corrosion or rusting
- Galvanising is an effective way of rust protection as it works even if the zinc coating becomes scratched or damaged
- The process can be done **electrolytically** or by dipping the metal parts into **baths** of **molten zinc**
- Zinc is also used to make an alloy called **brass**
- Brass contains 70% copper and 30% zinc
- The addition of zinc makes the alloy much **harder** and more **corrosion resistant** than copper alone

10.2.2 Uses of Metals

Uses of Aluminium, Copper & Mild Steel

Uses of Aluminium

USE	MOST IMPORTANT PROPERTY
AEROPLANE BODIES	HIGH STRENGTH-TO-WEIGHT RATIO (LOW DENSITY)
OVERHEAD POWER CABLES	GOOD CONDUCTOR OF ELECTRICITY
SAUCEPANS	GOOD CONDUCTOR OF HEAT
FOOD CANS	NON-TOXIC, RESISTANT TO CORROSION AND ACIDIC FOOD STUFFS
WINDOW FRAMES	RESISTANT TO CORROSION

Uses of Copper

USE	MOST IMPORTANT PROPERTY
ELECTRICAL WIRES	GOOD CONDUCTOR OF ELECTRICITY AND MALLEABLE
WATER PIPES	EASY TO WORK WITH AND BEND, NON-TOXIC AND UNREACTIVE (DOES NOT REACT WITH WATER)

Uses of Steel

TYPE OF STEEL	IRON ALLOYED WITH	USE	MOST IMPORTANT PROPERTY
MILD STEEL	0.25% CARBON	CAR BODY PANELS, WIRES	SOFT AND MALLEABLE
HIGH CARBON STEEL	0.5 – 1.4% CARBON	TOOLS AND CHISELS	HARD
LOW ALLOY STEEL	1 – 5% OF OTHER METALS (CR, NI, TI)	CONSTRUCTION, BRIDGES, HIGH SPEED TOOLS	HARD AND STRONG, LOW DUCTILITY AND MALLEABILITY
STAINLESS STEEL	20% CHROMIUM AND 10% NICKEL	CUTLERY AND SINKS, CHEMICAL PLANTS	STRONG AND RESISTANT TO CORROSION

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