Metals

The properties and uses of metals, alloys, the reactivity series and displacement reactions, corrosion and its prevention, and the extraction of iron in the blast furnace and aluminium by electrolysis.

9.1 Properties of metals

Property	Metals	Non-metals
thermal conductivity	good	poor (except graphite)
electrical conductivity	good	poor (graphite conducts)
malleable and ductile	yes (can be hammered and drawn into wires)	no, they are **brittle** when solid
melting and boiling points	usually **high**	usually **low**
appearance	shiny when polished	dull

Chemical properties

Metals form **positive ions** by losing electrons. The general reactions are:

Reaction with	Products	Example
dilute acids	salt + hydrogen	$\mathrm{Zn(s) + 2\,HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)}$
cold water	metal hydroxide + hydrogen	$\mathrm{2\,Na(s) + 2\,H_2O(l) \longrightarrow 2\,NaOH(aq) + H_2(g)}$
steam	metal oxide + hydrogen	$\mathrm{Mg(s) + H_2O(g) \longrightarrow MgO(s) + H_2(g)}$
oxygen	metal oxide	$\mathrm{2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)}$

How vigorously a metal reacts depends on its position in the **reactivity series** (section 9.4). Metals low in the series, such as copper, react with none of these under normal conditions.

9.2 Uses of metals

The use of a metal depends on its **physical properties**.

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Metal	Use	Property that makes it suitable
aluminium	making **aircraft**	**low density**, so the aircraft is light
aluminium	**overhead electrical cables**	low density and good electrical conductivity (a steel core adds strength)
aluminium	**food containers**	**resistant to corrosion** (because of its oxide layer)
copper	**electrical wiring**	good electrical **conductivity** and it is **ductile**, so it can be drawn into wires

9.3 Alloys and their properties

> **An alloy is a mixture of a metal with other elements (usually other metals).**

Alloy	Made of	Use
brass	copper and zinc	musical instruments, door handles, taps
stainless steel	iron with other elements such as chromium, nickel and carbon	cutlery, kitchen sinks, surgical instruments

Alloys are **harder and stronger** than the pure metals. In a **pure metal** the atoms are all the **same size**, in regular layers, and the layers can **slide over each other** easily, so the metal is soft and bends. In an **alloy**, the atoms (or ions) of the other element are a **different size**, which **disrupts the regular layers**, so the layers **can no longer slide** over each other as easily.

Stainless steel is used for **cutlery** because it is **hard** and **resists rusting**.

NOTE

To identify an alloy from a diagram, look for **atoms of different sizes** mixed in the lattice. A pure metal has only one size of atom, in a regular pattern.

9.4 Reactivity series

The **reactivity series** lists metals in order of their **tendency to form positive ions** (to lose electrons).

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Order	Metal	Reaction with cold water	Reaction with dilute hydrochloric acid
most reactive	**potassium** (K)	very vigorous	explosive
	sodium (Na)	vigorous	explosive
	calcium (Ca)	steady, fizzing	very vigorous
	magnesium (Mg)	very slow in cold water, reacts with **steam**	vigorous
	aluminium (Al)	protected by an oxide layer	slowly at first, then vigorous
	carbon (C)		
	zinc (Zn)	reacts with steam only	steady
	iron (Fe)	reacts with steam only	slow
	hydrogen (H)		
	copper (Cu)	no reaction	no reaction
	silver (Ag)	no reaction	no reaction
least reactive	**gold** (Au)	no reaction	no reaction

A way to remember the order: **P**lease **S**top **C**rying **M**y **A**unt **C**arol **Z**ebra **I**s **H**ungry **C**ats **S**hould **G**o. (Carbon and hydrogen are included for comparison; they are not metals in this sense.)

Reactions with water and steam. Potassium, sodium and calcium react with **cold water** to give the hydroxide and hydrogen. Magnesium reacts with **steam** to give magnesium oxide and hydrogen. The unreactive metals below do not react.

$$\mathrm{Ca(s) + 2\,H_2O(l) \longrightarrow Ca(OH)_2(aq) + H_2(g)}$$

Reactions with dilute hydrochloric acid. Magnesium, zinc and iron react to give a salt and hydrogen (**magnesium most vigorously**), but **copper, silver and gold do not react**, because they are below hydrogen in the series.

Displacement reactions

> **A more reactive metal displaces a less reactive metal from a solution of its salt.**

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The more reactive metal forms its **positive ion** (loses electrons) and the less reactive metal ion is **reduced** to the metal.

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

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

Metal added to solution of	Mg^{2+}	Zn^{2+}	Fe^{2+}	Cu^{2+}	Ag^{+}
magnesium	–	displaces	displaces	displaces	displaces
zinc	no reaction	–	displaces	displaces	displaces
iron	no reaction	no reaction	–	displaces	displaces
copper	no reaction	no reaction	no reaction	–	displaces
silver	no reaction	no reaction	no reaction	no reaction	–

The blue colour of copper(II) sulfate solution fades when a more reactive metal displaces copper, and a pink-brown solid of copper forms.

> **TRY IT**
>
> A metal X displaces copper from copper(II) sulfate solution, but does not displace zinc from zinc sulfate solution. Where is X in the reactivity series?

X is **more reactive than copper** and **less reactive than zinc**, so it is between zinc and copper. (Iron is an example.)

> **TRY IT**
>
> Three metals P, Q and R are added to dilute acid. Q gives fast bubbles, P gives slow bubbles and R gives no bubbles. Put them in order of reactivity.

The more bubbles, the more reactive the metal, so the order is Q **(most), then** P**, then** R **(least)**.

The apparent unreactivity of aluminium

Aluminium is high in the reactivity series, but it seems unreactive. This is because it is covered with a thin, strong layer of **aluminium oxide** that stops the acid or water reaching the metal. If the layer is removed, aluminium reacts vigorously.

Note Academy

9.5 Corrosion of metals

Rusting is the corrosion of **iron and steel**.

> **Iron rusts when it is in contact with both oxygen (air) and water. Rust is hydrated iron(III) oxide,** $\mathrm{Fe_2O_3 \cdot \mathit{x}\, H_2O}$**.**

The experiment shows both are needed: a nail in air and water **rusts**; a nail in boiled water (no dissolved air) covered with oil **does not rust**, and a nail in dry air (with calcium chloride to absorb water vapour) **does not rust**.

Preventing rust

Barrier methods work by **excluding oxygen or water**:

- **painting** (cars, bridges, ships),
- **greasing or oiling** (machine parts, bicycle chains),
- **coating with plastic** (fridge shelves, wire fences),
- coating with another metal such as tin or chromium.

Sacrificial protection. A **more reactive metal** such as zinc or magnesium is attached to the iron. The more reactive metal **loses electrons** in preference to the iron, so it corrodes instead of the iron:

$$\mathrm{Zn \longrightarrow Zn^{2+} + 2\,e^-}$$

The electrons flow to the iron, which keeps it from forming $\mathrm{Fe^{2+}}$ ions. The zinc or magnesium is **used up** and must be replaced. Blocks of zinc or magnesium are bolted to the hulls of ships and to underground pipes.

Galvanising is coating iron with **zinc**. It protects by **both** methods: the zinc layer is a **barrier**, and if it is scratched, the zinc still protects the iron by **sacrificial protection**, because zinc is more reactive than iron.

> NOTE
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> Tin plating is only a **barrier**: if the layer is scratched, the iron rusts **faster**, because iron is more reactive than tin. Zinc is different, because it is more reactive than iron.

9.6 Extraction of metals

A metal is found in the Earth as an **ore**, a rock that contains the metal compound in a large enough amount to be worth extracting. The method depends on the metal's position in the reactivity series.

Note Academy

Position	Method	Examples
very unreactive (gold, silver)	found **uncombined** as the metal in nature	gold, silver
less reactive than carbon	**reduction** of the ore (oxide) with **carbon** (heating with coke)	iron, zinc, copper
more reactive than carbon	**electrolysis** of the molten compound (carbon cannot reduce them)	aluminium, sodium, magnesium

The more reactive the metal, the **more difficult** it is to extract, as its compounds are **more stable**.

Iron: the blast furnace

The raw materials are **iron ore** (hematite, $\mathrm{Fe_2O_3}$), **coke** (carbon), **limestone** ($\mathrm{CaCO_3}$) and **hot air**.

1. **Coke burns** in the hot air, giving carbon dioxide and heat energy:

$$\mathrm{C(s) + O_2(g) \longrightarrow CO_2(g)}$$

2. The **carbon dioxide is reduced** to carbon monoxide by more hot coke:

$$\mathrm{CO_2(g) + C(s) \longrightarrow 2\,CO(g)}$$

3. **Carbon monoxide reduces iron(III) oxide** to iron. The iron(III) oxide loses oxygen, so it is reduced:

$$\mathrm{Fe_2O_3(s) + 3\,CO(g) \longrightarrow 2\,Fe(l) + 3\,CO_2(g)}$$

4. **Limestone** removes impurities. It decomposes with heat (thermal decomposition) to give calcium oxide:

$$\mathrm{CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)}$$

5. The basic calcium oxide reacts with acidic impurity silicon(IV) oxide, $\mathrm{SiO_2}$, to form **slag**, calcium silicate:

$$\mathrm{CaO(s) + SiO_2(s) \longrightarrow CaSiO_3(l)}$$

The **molten iron** is denser, so it collects at the **bottom**, and the **slag floats on top** of it. Both are tapped off separately. The slag is used for road building.

Aluminium: electrolysis

Aluminium is extracted from **aluminium oxide**, $\mathrm{Al_2O_3}$, purified from the ore **bauxite**. It is more reactive than carbon, so it is made by **electrolysis**.

- Aluminium oxide has a very high melting point ($2050\,^\circ\mathrm{C}$), so it is **dissolved in molten cryolite**. This **lowers the temperature** of the electrolyte to about $950\,^\circ\mathrm{C}$ and **saves energy**.

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- The electrodes are **carbon** (graphite).
- At the **cathode** (negative), aluminium ions are reduced: $\mathrm{Al^{3+} + 3\,e^- \longrightarrow Al}$. Molten aluminium sinks to the bottom and is tapped off.
- At the **anode** (positive), oxide ions are oxidised: $\mathrm{2\,O^{2-} \longrightarrow O_2 + 4\,e^-}$.

The **oxygen** formed at the anode reacts with the hot **carbon anodes**, so they **burn away** to carbon dioxide and must be **replaced regularly**:

$$\mathrm{C\,(s) + O_2\,(g) \longrightarrow CO_2\,(g)}$$

> **TRY IT**
>
> Explain why aluminium cannot be extracted from its oxide by heating with carbon.

Aluminium is **more reactive than carbon**, so carbon cannot take oxygen away from aluminium oxide. Electrolysis is needed.

Note Academy