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TOPIC 9: METALS

THE ABOUT

CHAPTER ANALYSIS

TIME

- Heavy content chapter
- 5 **key** concepts
- 2 **advanced** concepts

EXAM

- Always tested in exams, MCQ and FRQ
- Require a little knowledge from chapters like:
  → Periodic Table, Oxidation & Reduction

WEIGHTAGE

- Heavy_Medium overall weightage
- Constitute to **5.5%** of marks for past 5 year papers

KEY CONCEPT

METALS
PHYSICAL PROPERTIES OF METAL ALLOYS

PHYSICAL PROPERTIES OF METAL

PHYSICAL PROPERTIES OF METAL

1) Metals are **ductile** (can be stretched to form wires).

2) Metals are **malleable** (shaped into different shapes by applying pressure).

3) Metals are all good electrical conductors and good heat conductors.

4) Metals have **high melting and boiling points** and exist as solids at room temperature.

5) Metals have **high density** (due to the closely packed arrangement of atoms).

6) Metals are **strong and shiny**.

Exceptions:

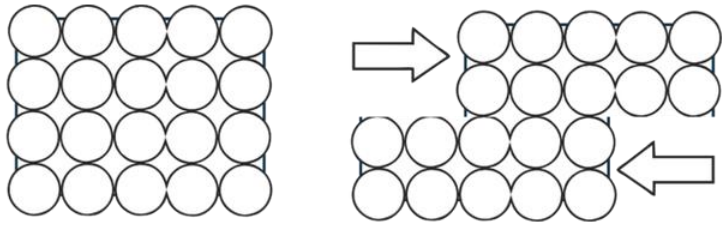
- Mercury has a low melting point (–39 °C) and exists as a liquid at room temperature.

- Group I metals such as lithium, sodium and potassium have low densities, allowing them to float on water.

ALLOYS

PURE METAL

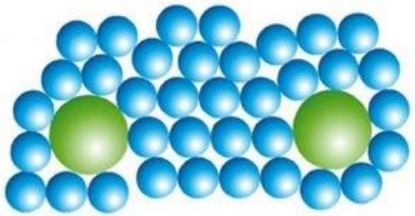
In pure metals, the **layers of atoms** are able to **slide over one another easily**, making them too soft and weak for most practical uses.



ALLOYS

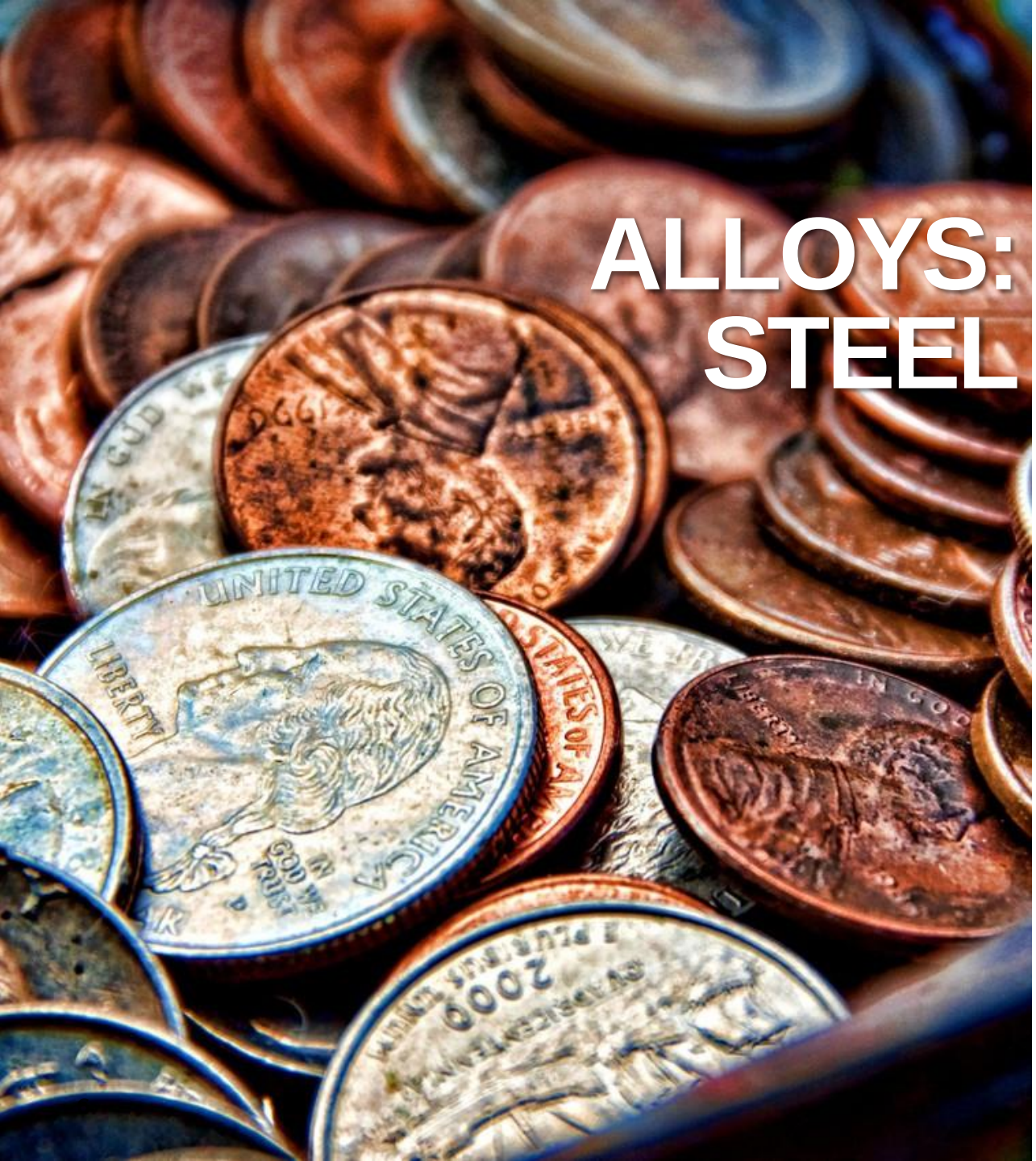
Alloys are **mixtures of metals with other elements.**

In alloys, elements with different sizes distort the orderly arrangement of the metal atoms and thus cannot slide over as easily. This makes alloys much **stronger and harder.**



Examples:

Steel: Iron, carbon (bodies of cars)
Brass: Copper, zinc (electrical plugs)
Bronze: Copper, tin (trophies)
Stainless Steel: iron, carbon, chromium, nickel (cutlery, medical instruments)



ALLOYS: STEEL

STEEL

Steel is an alloy which is a mixture of iron with carbon or other metals.

Controlling the composition of steel will result in **high carbon** steels **& low carbon** steels**.**

Category	Type of Steel	Uses	Special Properties
Carbon Steels	**Mild Steel** *0.25% Carbon*	Car bodies and machinery	**Hard, strong and malleable**
	High Carbon Steel *0.45 – 1.5% Carbon*	Cutting and boring tools, e.g. knives, hammers	**Strong but brittle** (more carbon atoms to prevent sliding)
Alloy Steels	**Stainless Steel** *Alloy of iron, chromium, nickel & carbon.*	Equipments in chemical plants, cutlery, surgical instruments	Extremely **durable, resistant to rust and corrosion** even when heated

Qn: Explain how the properties of low carbon and high carbon steel differ.

Low carbon steel is softer as it is more malleable.

High carbon steel contains more carbon atoms which prevent sliding of the iron atoms. Hence, high carbon steel is harder but brittle.

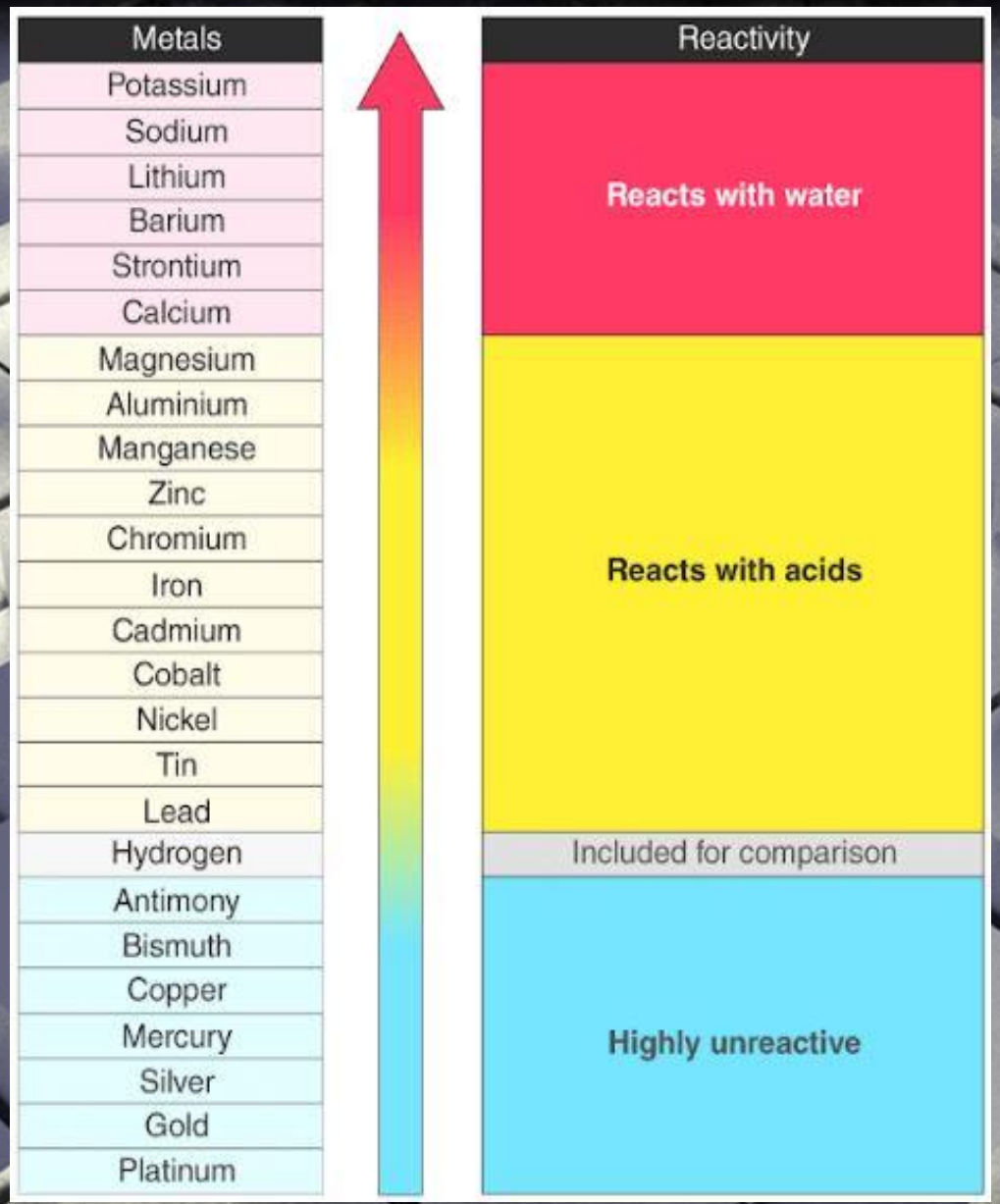
KEY CONCEPT

REACTIVITY SERIES
CHEMICAL REACTIONS OF METALS
DISPLACEMENT, DECOMPOSITION, RUSTING

MUST KNOW

Complete Summary Table

Acroymn	Metal	Periodic Table	Stability	Reaction with water	Reaction with acid
Please	Potassium (**K**)	Group I	Compound broken down by **electrolysis**	Can react with **cold water** to form metal hydroxide	React with **acid**
Stop	Sodium (**Na**)				
Calling	Calcium (**Ca**)	Group II			
Me	Magnesium (**Mg**)			Can react with **steam** to form metal oxide	
A	Aluminium (**Al**)	Group III			
Cute	Carbon (**C**)	---	---		
Zombie	Zinc (**Zn**)	Transition Metals	Compound broken down by **reduction with carbon**		
I	Iron (**Fe**)				
Like	Lead (**Pb**)			Does not react with steam or cold water	
Hwa	Hydrogen (**H**)	---	---		---
Chong	Copper (**Cu**)	Unreactive Metals	Compound broken down by **thermal decomposition**		Does not react with acid
Sexy	Silver (**Ag**)				
Guys/Girls	Gold (**Au**)				



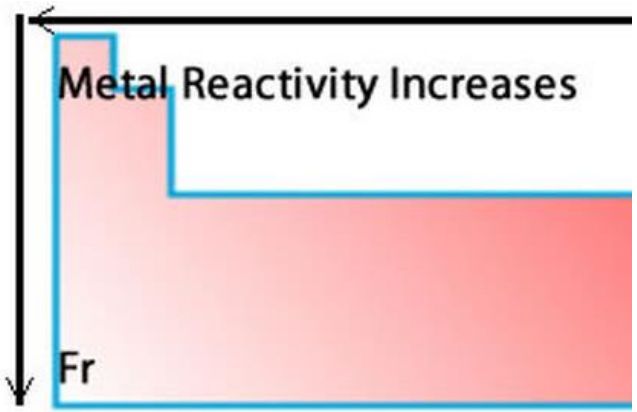
REACTIVITY OF METALS

Reactivity **increases down the group** and from **right to left**.

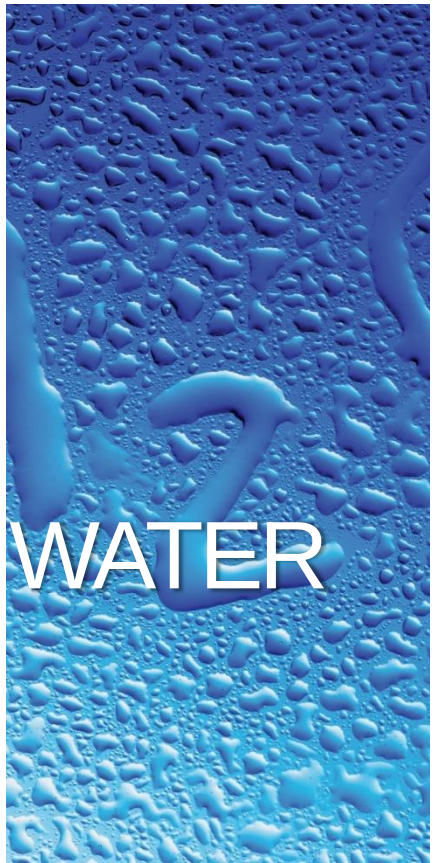
In other words, **Group I metals are the strongest**.

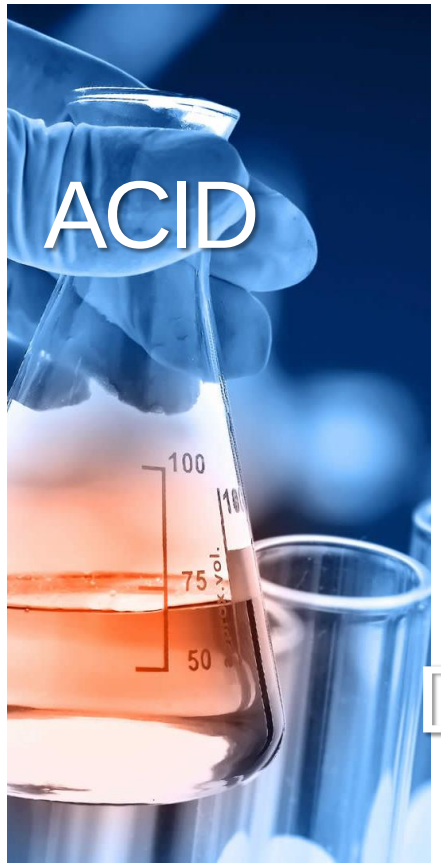
The **more valence shells** an metal has, the **more readily it can lose its valence electrons**, hence it is more reactive.

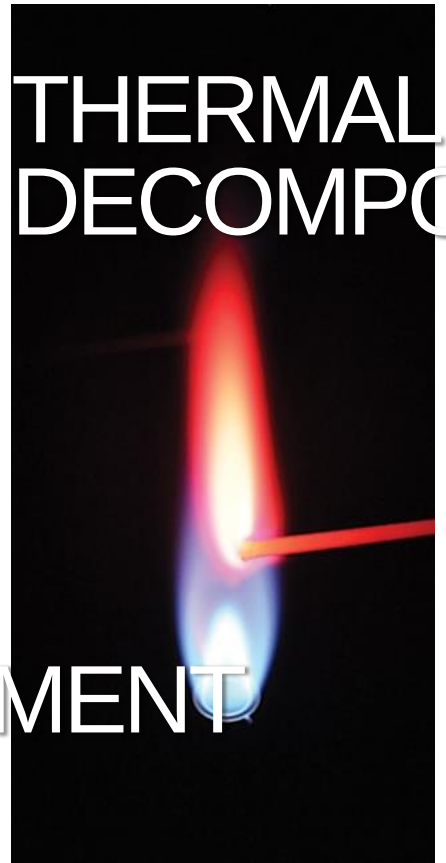
(from chapter 'Periodic Table'.)

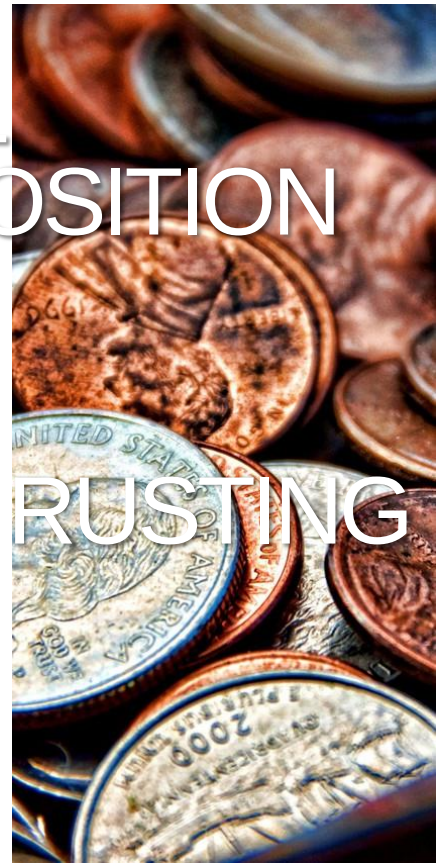


CHEMICAL REACTIONS OF METALS









METAL + WATER

METAL + WATER → METAL OXIDE / HYDROXIDE + HYDROGEN GAS

When metals react with water, metal oxide or hydroxide is formed, along with hydrogen gas*.

Reactive metals are able to react with cold H_2O.

Less reactive metals are only able to react with steam.

Unreactive metals cannot react with water at all.

*Test for hydrogen gas using lighted splint, it should extinguish with 'pop sound.

METAL + WATER

Metal	Speed of Reaction	Observation	Chemical Equation
Potassium (**K**)	**explosively** in cold water	burns with **lilac flame**	$2K (s) + 2H_2O (l) \rightarrow 2KOH (aq) + H_2 (g)$
Sodium (**Na**)	**violently** in cold water	burns with **yellow flame**	$2Na (s) + 2H_2O (l) \rightarrow 2NaOH (aq) + H_2 (g)$
Calcium (**Ca**)	**readily** in cold water	**vigorous effervescence**	$Ca (s) + 2H_2O (l) \rightarrow Ca(OH)_2 (aq) + H_2 (g)$
Magnesium (**Mg**)	very **slowly** in cold water **violently** with steam	**little effervescence** burns with white glow	$Mg (s) + H_2O (g) \rightarrow MgO (s) + H_2 (g)$
Aluminium (**Al**)	**readily** in steam		$2Al (s) + 3H_2O (g) \rightarrow Al_2O_3 (s) + 3 H_2 (g)$
Zinc (**Zn**)	**readily** in steam	ZnO is **yellow** when **hot** **white** when **cooled**	$Zn (s) + H_2O (g) \rightarrow ZnO (s) + H_2 (g)$
Iron (**Fe**)	**slowly** in steam	requires **constant heating**	$3Fe (s) + 4H_2O (g) \rightarrow Fe_3O_4 (s) + 4H_2 (g)$

METAL + ACID

METAL + ACID → SALT + HYDROGEN GAS

When metals react with acid, salt and hydrogen gas* is produced.

More reactive metals cause more vigorous reactions.

Less reactive metals cause less effervescence of hydrogen gas.

Less reactive metals only react with warm dilute hydrochloric acid.

*Test for hydrogen gas using lighted splint, it should extinguish with 'pop sound.

METAL + ACID

Metal	Speed of Reaction	Chemical Equation
Potassium (**K**)	**Explosively** in acid	$2\ K(s) + 2\ HCl\ (aq) \rightarrow 2KCl\ (aq) + H_2\ (g)$
Sodium (**Na**)	**Explosively** in acid	$2\ Na(s) + 2HCl\ (aq) \rightarrow 2NaCl\ (aq) + H_2\ (g)$
Calcium (**Ca**)	**Violently** in acid	$Ca(s) + 2HCl\ (aq) \rightarrow CaCl_2\ (aq) + H_2\ (g)$
Magnesium (**Mg**)	**Rapidly** in acid	$Mg(s) + 2HCl\ (aq) \rightarrow MgCl_2\ (aq) + H_2\ (g)$
Aluminium (**Al**)	**Readily** in acid	$4\ Al\ (s) + 6HCl\ (aq) \rightarrow 2\ Al_2Cl_3\ (s) + 3\ H_2\ (g)$
Zinc (**Zn**)	**Moderately** in acid	$Zn(s) + 2HCl\ (aq) \rightarrow ZnCl_2\ (aq) + H_2\ (g)$
Iron (**Fe**)	**Slowly** in acid	$Fe(s) + 2\ HCl\ (aq) \rightarrow FeCl_2\ (aq) + H_2\ (g)$
Lead (**Pb**)	**Slowly** in acid	$Pb(s) + 2\ HCl\ (aq) \xrightarrow{\text{heating}} PbCl_2\ (s) + H_2\ (g)$

DISPLACEMENT

Displacement reactions occur when a **more reactive metal displaces a less reactive metal** from its solution or oxide.

This is due to the former being able to **lose its electrons more readily** to form cations.

We can use displacement reaction to find out relative reactivity of two metals.

If a metal precipitates when another metal is added to the solution, then the added metal is the more reactive one.

If no change is observed (i.e. no reaction) then the metal added is less reactive.

TAKE NOTE

Carbon (non-metal) can also displace less reactive metals from their oxides.

For example, carbon can remove copper from copper(II) oxide:

$$2\ CuO\ (s) + C\ (s) \rightarrow 2\ Cu\ (s) + CO_2\ (g)$$

DISPLACEMENT OF METALS FROM SOLUTION

EXAMPLE

When magnesium ribbon is added to blue copper(II) sulphate solution, some of the magnesium ribbon gets eroded while a pinkish-brown solid forms on the surface of the magnesium ribbon and the blue solution decolourises.

Equation:

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

Explanation:
Copper is displaced, forming a pinkish-brown solid.
Magnesium displaced copper from its solution, its Mg^{2+} ions are colourless, hence solution decolourises.

DISPLACEMENT OF METALS FROM METAL OXIDE

EXAMPLE

Aluminium powder is mixed with iron(III) oxide and ignited with a fuse. Aluminium displaces iron from its oxide to form aluminium oxide:

$$\mathbf{2\ Al(s) + Fe_2O_3\ (s) \rightarrow Al_2O_3\ (s) + 2\ Fe\ (l)}$$

Explanation:
Aluminium donates its electrons to iron(III) ions in the oxide because it can lose electrons more readily than iron.

Therefore, it can be concluded that aluminium is more reactive than iron.

THERMAL DECOMPOSITION

The **more reactive a metal, the more heat-stable** its carbonate as it has a higher tendency to remain as a cation.

More reactive metal's carbonates require **electrolysis** while the less reactive ones can be broken down by **reduction with carbon** or by **heating**.

Metal	Periodic Table	Stability
Potassium (**K**)	Group I	Compound broken down by **electrolysis**
Sodium (**Na**)		
Calcium (**Ca**)	Group II	
Magnesium (**Mg**)		
Aluminium (**Al**)	Group III	
Carbon (**C**)	---	---
Zinc (**Zn**)	Transition Metals	Compound broken down by **reduction with carbon**
Iron (**Fe**)		
Lead (**Pb**)		
Hydrogen (**H**)	---	---
Copper (**Cu**)	Unreactive Metals	Compound broken down by **thermal decomposition**
Silver (**Ag**)		
Gold (**Au**)		

THERMAL DECOMPOSITION OF METAL CARBONATES

Metal carbonates decompose when heated strongly, producing the metal oxide and carbon dioxide gas.

Example:

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

Explanation:

The more reactive a metal, the more heat-stable is its carbonate as it has a higher tendency to remain as a cation.

Therefore, more energy is required to thermally decompose the carbonate of a more reactive metal than that of a less reactive one.

*Carbon is more reactive than these few metals. So carbon can "displace" these metals from their carbonate. This is the reason why we memorise carbon as part of the reactivity series.

Similarly, anything below hydrogen will not be able to displace it, hence those metals below hydrogen are not able to react with acid.

RUSTING

The **corrosion of iron and steel** is called rusting. This occurs when iron corrodes due to a chemical attack by air and water.

Rust is a brown solid. It is **hydrated iron(III) oxide** with the chemical formula:

$$Fe_2O_3 . xH_2O$$

Iron must be in contact with both air (oxygen) and water to rust.

Other factors can speed up rusting, such as dissolved salt. Seawater will cause rusting faster due to the presence of ions in seawater that act as a charge carrier.

During rusting, electrons are transferred from the iron metal to the oxygen and water molecules.

The presence of ions facilitate this transfer of electrons, speeding up the process of rusting.

PREVENTING RUSTING

Surface Protection

Surface or barrier protection prevents rusting by covering the surface of the object with a layer of substance.

This stops air and water from coming into contact with iron (or steel) under the protection.

Paint, oil, plastic and metal plating are commonly used to protect the surfaces of iron and steel objects.

Sacrificial Metals

If iron is covered by a **more reactive metal** like magnesium or zinc, then the rusting of iron is greatly minimised. These reactive metals react in place of iron.

However, the sacrificed metal has to be replaced before it is all used up.

Stainless Steel

Stainless steel is an **iron alloy** which also contains chromium or nickel.

It does not rust easily, because these metals react with the oxygen in the air to produce a **stable oxide layer**, which acts as a barrier to prevent the iron within from coming into contact with air.

KEY CONCEPT

EXTRACTION OF METALS
ELECTROLYSIS, REDUCTION, HYDROGEN
BLAST FURNACE

EXTRACTION OF METALS

The method used to extract a given metal from its ore depends upon the **reactivity of the metal** and the **stability of the ore**.

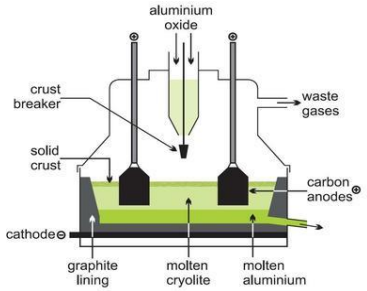
In general, very reactive metals are extracted using electricity, while less reactive metals are extracted by reduction.

Metal	Extraction method	Reduction by hydrogen
Potassium (**K**)	**Electrolysis**	**Cannot be reduced by hydrogen**
Sodium (**Na**)		
Calcium (**Ca**)		
Magnesium (**Mg**)		
Aluminium (**Al**)		
Carbon (**C**)	---	
Zinc (**Zn**)	**Displacement / reduction with carbon**	
Iron (**Fe**)		
Lead (**Pb**)		
Hydrogen (**H**)	---	**Reduced by hydrogen**
Copper (**Cu**)	**Heating in air**	
Silver (**Ag**)	**Exist naturally as metal**	
Gold (**Au**)		

REDUCTION BY ELECTROLYSIS

Electrolysis is the most powerful extraction method. As it uses a lot of electricity, it is a highly expensive process.

Hence, electrolysis is only used for the **most reactive metals** like potassium, sodium, calcium, magnesium and aluminium.



REDUCTION BY CARBON

Metal oxides that can be reduced by carbon are **zinc, iron, tin and lead.**

Lead(II) oxide is reduced to pure lead metal by carbon.

$$PbO\ (s) + C\ (s) \rightarrow Pb\ (s) + CO\ (g)$$

Zinc oxide is reduced to pure zinc metal by carbon monoxide.

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

There is an overlap!

REDUCTION BY HYDROGEN

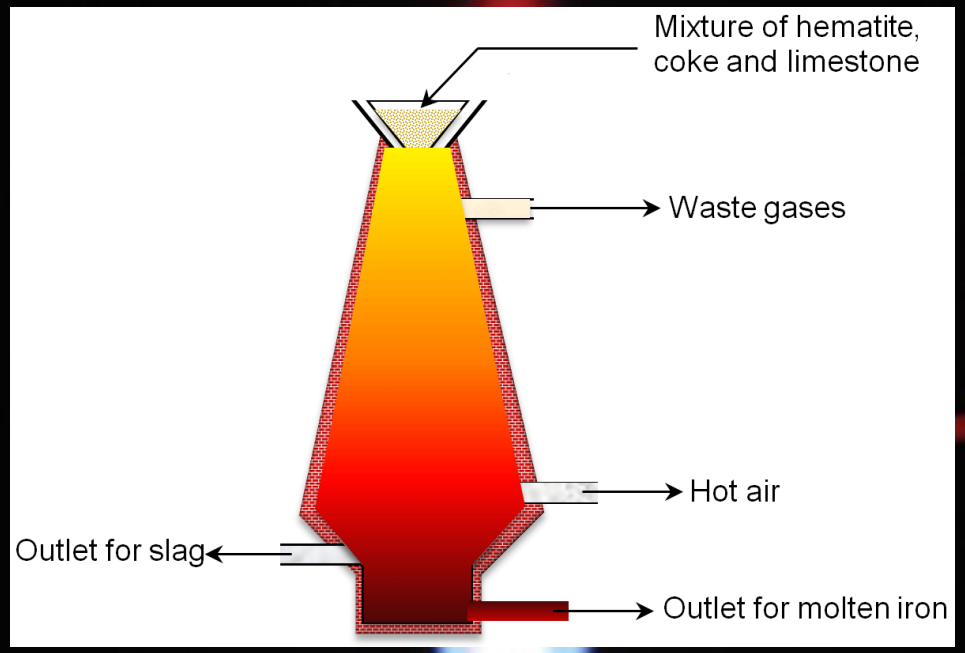
Metal oxides that can be reduced by hydrogen are **iron, copper and lead.**

$$CuO\ (s) + H_2\ (g) \rightarrow Cu\ (s) + H_2O\ (g)$$
$$PbO\ (s) + H_2\ (g) \rightarrow Pb\ (s) + H_2O\ (g)$$

*Even though **iron is above hydrogen** in the reactivity series, **iron(II) oxide can be reduced by hydrogen** to form pure iron metal and steam.

$$FeO\ (s) + H_2\ (g) \rightarrow Fe\ (s) + H_2O\ (g)$$

EXTRACTION OF IRON



BLAST FURNACE

These are the following reactions in a blast furnace:

Production of carbon dioxide

Carbon in coke burns in hot air to produce carbon dioxide.

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

Limestone decomposed by heat to form carbon dioxide and calcium oxide.

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

Production of carbon monoxide

As carbon dioxide rises up in the blast furnace, it reacts with more coke to form carbon monoxide.

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

The first 2 steps are meant to produce CO that will reduce the iron (III) oxide!

Reduction of haematite to iron

Carbon monoxide reduces iron(III) oxide in haematite to iron.

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

Molten iron is formed and it runs to the bottom of the blast furnace.

Waste gases such as nitrogen, carbon dioxide and unreacted carbon monoxide escape from the top of the blast furnace.

Removal of impurities

Impurities such as silicon(IV) oxide are removed by calcium oxide.

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

$CaSiO_3$ is called calcium silicate or slag. It floats on top of molten iron. It is removed separately from the iron through another tap.

RECYCLING

RECYCLING OF METALS

Metals are finite resources and need to be conserved.

The amount of metal ores in the Earth is limited. If metal extraction continues at present rates, the supplies of many metals will run out.

Hence, there is a need to **recycle** metals.

	Upside	Downside
Economic	Save the costs of extracting new metals from their ores. Fewer landfills to dispose used metal objects.	Recycling is very expensive, like collecting, transporting and separating the scrap metals.
Social	Conserve limited amount of metals on Earth. More land will be available if ore mining is reduced.	Time and effort for recycling as a way of life.
Environmental	Reduces waste gases including carbon monoxide which are formed in the blast furnace)	The recycling process may cause pollution if not done properly (e.g. metal fumes from the recycling process)

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