2020
e-Book
Edition

METALS CRACKED

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CRACK BOX

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This author of this document is Trishit Dedhia.

Special thanks to Mandira Bhattacharyya, IGCSE Chemistry teacher at Don Bosco International School.

If you have any doubts, queries or feedback, you may reach the CrackBox Team using email: crackboxofficial@gmail.com.

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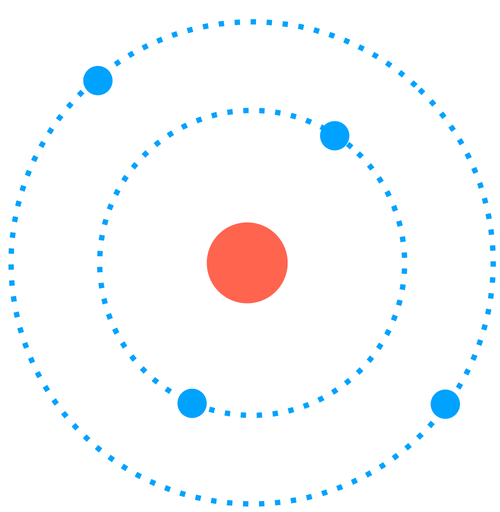
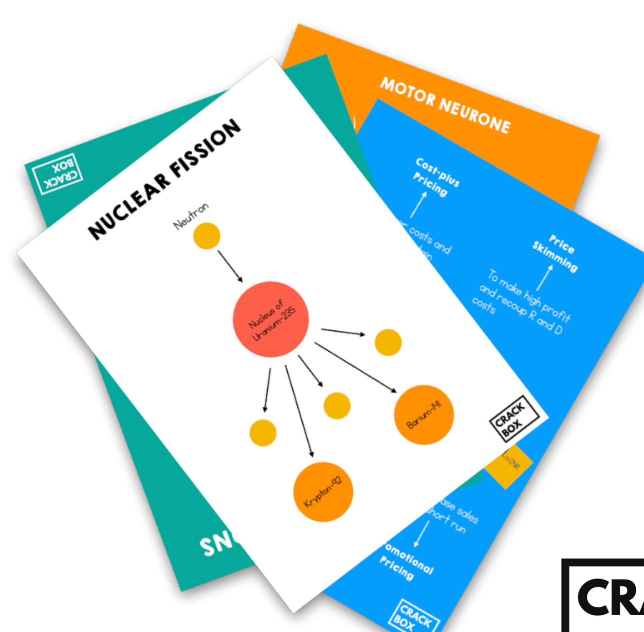
Textbooks *Remastered*

Pre-highlighted key terms and phrases that are important!



Illustrations that help visualise and understand concepts better!

The book gives you alerts whenever a poster or game is available!



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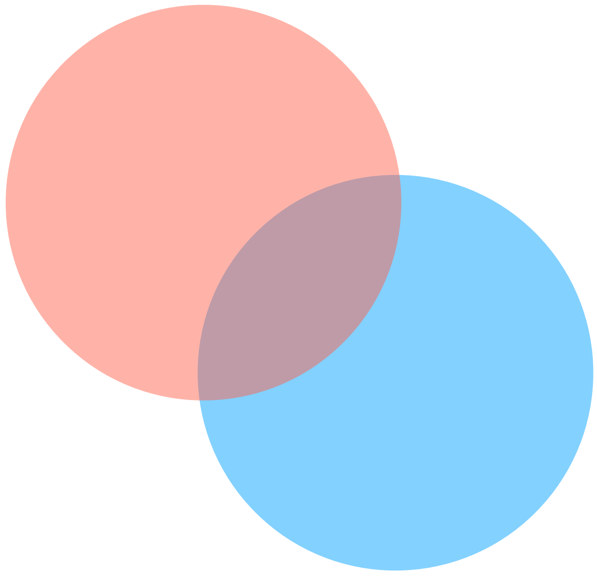
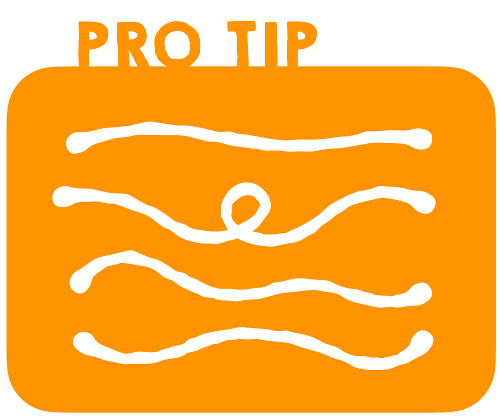


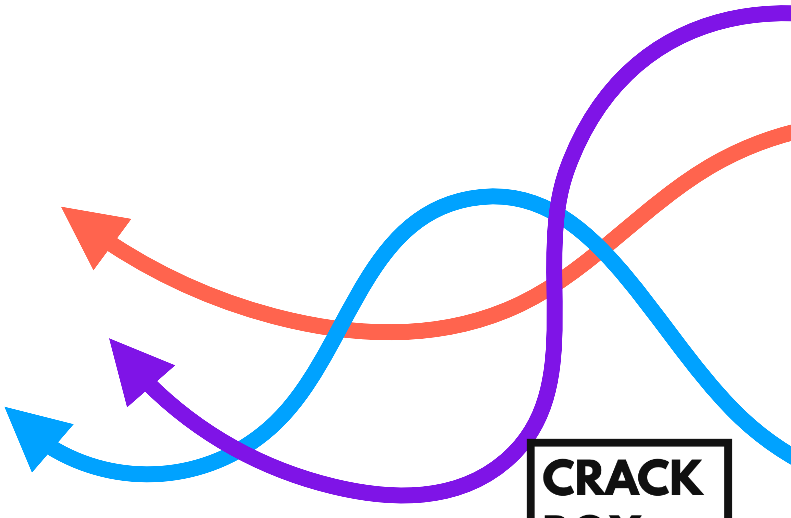
Unclustered, concise, to the point, organised and need to know information.

By the students, for the students. Learn like you learn from your peers!



Includes pro-tips for student insights and techniques!

Arrows and multiple colours to make the reading experience much better!



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SUBCONTENTS

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Reactivity Series

As we know, metals have similar physical properties as most of them are:

SONOROUS

MALLEABLE

LUSTROUS

But most of them are chemically different, as some react more vigorously and some just do not.

IRON RUSTS WHEN LEFT UNPROTECTED

GOLD, UNLIKE IRON DOESN'T RUST WHEN EXPOSED

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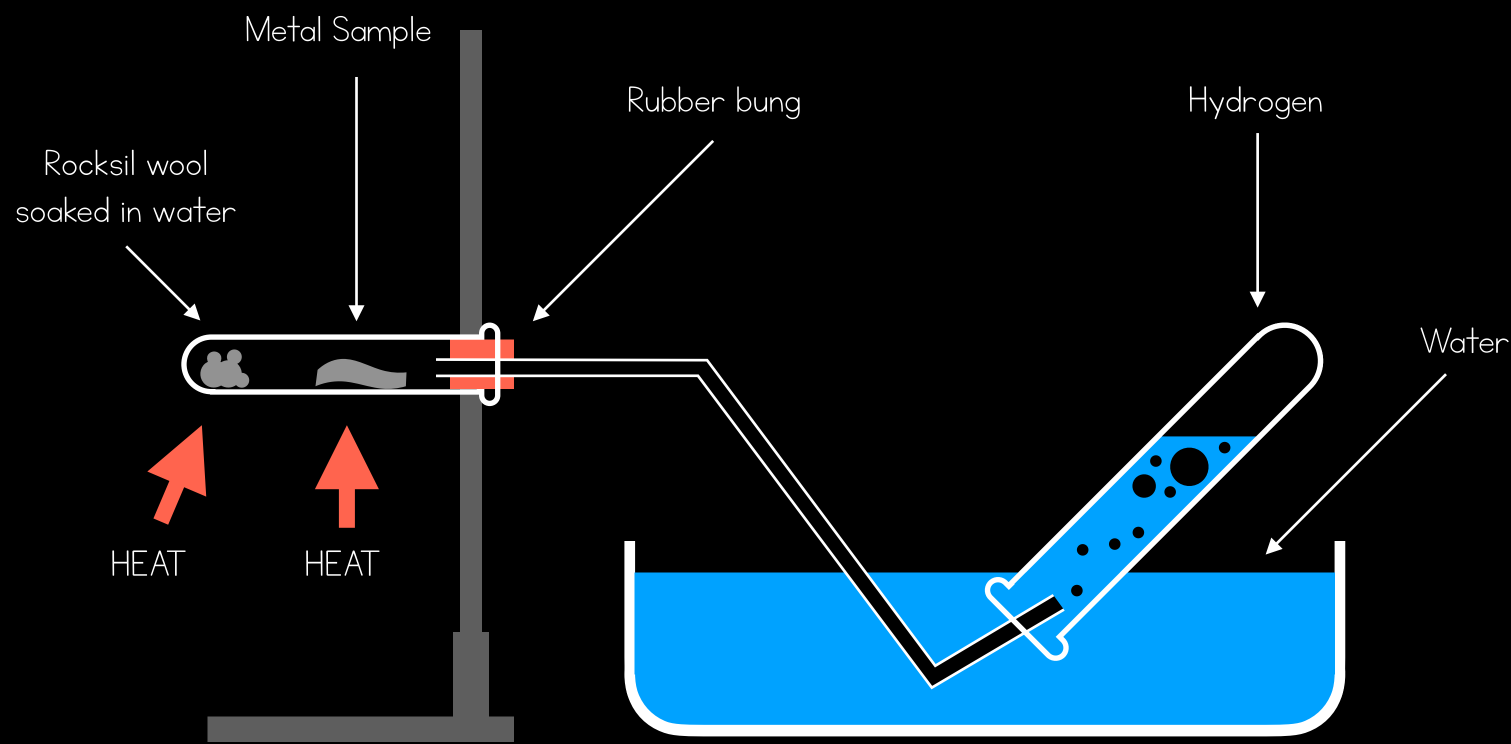
REACTIVITY SERIES	REACTION WITH DILUTE ACID	REACTION WITH AIR/ OXYGEN	REACTION WITH WATER	EASE OF EXTRACTION
K	Produce H2 with decreasing vigour	Burn very brightly and vigorously	Produce H2 with decreasing vigour with cold water	Difficult
Na				
Ca		Burn to form an oxide with decreasing vigour	React with steam with decreasing vigour	Easier
Mg				
Al				
C				
Zn				
Fe				
Pb		React slowly to form the oxide	Don't react with cold water or steam	
H				
Cu	Don't react with dilute acids			
Ag		Don't react		Found Native
Au				
Pt				

How a metal is made to react with steam:



Some balanced equation examples:

$$Mg(s) + 2HCl(aq) \xrightarrow{HEAT} MgCl_2(aq) + H_2(g)$$

$$2Mg(s) + O_2(g) \xrightarrow{HEAT} MgO(s)$$

$$2Na(s) + 2H_2O(l) \xrightarrow{HEAT} 2NaOH(aq) + H_2(g)$$

$$Mg(s) + H_2O(g) \xrightarrow{HEAT} MgO(s) + H_2(g)$$

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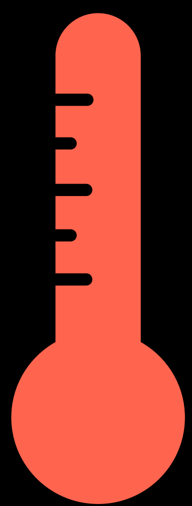
Thermal Decomposition

As we know, metals can also be present as:

- Metal Nitrates
- Metal Carbonates
- Metal Hydroxides
- Metal Oxides

But these metals can be detached in the presence of heat and that process is know as Thermal Decomposition!



Before we begin with each type, we must classify each metal in the reactivity series.

Potassium
Sodium

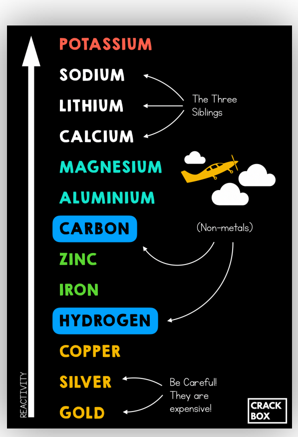
Reactive Metals

Calcium
Magnesium
Aluminium
Zinc
Iron
Lead
Copper

Moderately Reactive Metals

Silver
Platinum
Gold

Unreactive Metals



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SUBTOPIC

Thermal Decomposition of Metal Nitrates

When nitrates of reactive metals are heated, they decompose to produce the metal nitrite and oxygen.

$$2NaNO_3(s) \xrightarrow{\text{HEAT}} 2NaNO_2(s) + O_2(g)$$

NOTE:

NITRATE is NO_3^-

NITRITE is NO_2^-

When nitrates of moderately reactive metals are heated, they produce BROWN fumes of nitrogen dioxide as well as the metal oxide and oxygen.

$$2Mg(NO_3)_2(s) \xrightarrow{\text{HEAT}} 2MgO(s) + 4NO_2(g) + O_2(g)$$

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When nitrates of Unreactive Metals are heated, they decompose to produce the metal, nitrogen oxide and oxygen.

$$2AgNO_3(s) \xrightarrow{\text{HEAT}} 2Ag(s) + 2NO_2(g) + O_2(g)$$

Reactive

Form metal nitrite and oxygen

Moderate

Form metal oxide, nitrogen dioxide and oxygen

Unreactive

Form metal, nitrogen dioxide and oxygen

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SUBTOPIC

Thermal Decomposition of Metal Carbonates

All metal carbonates that exist undergo thermal decomposition to give the metal oxide and carbon dioxide.

$$\text{Metal Carbonate} \xrightarrow{\text{HEAT}} \text{Metal Oxide} + \text{Carbon Dioxide}$$

EXAMPLE

$$CaCO_3 \longrightarrow CaO + CO_2$$

More reactive the metal is, more heat is required to decompose its carbonate form!

Reactive & Moderate

Form metal oxide and carbon dioxide

Unreactive

Too Unstable so Don't exist

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SUBTOPIC

Thermal Decomposition of Metal Oxides

Theoretically, all metal oxides can be thermally decomposed, but reactive metals and some moderately reactive metals like aluminium require so much of energy that it is impractical to decompose them.

However, it's possible to decompose oxides of unreactive metals like silver.

$$2Ag_2O(s) \xrightarrow{\text{HEAT}} 4Ag(s) + O_2(g)$$

Reactive

Possible but impractical

Moderate

Possible but only some are practical

Unreactive

Form the metal and oxygen

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Thermal Decomposition of Metal Hydroxides

Hydroxides of moderately reactive metals decomposed form the metal oxide and water when heated.

$$Ca(OH)_2(s) \xrightarrow{\text{HEAT}} CaO(s) + H_2O(g)$$

This process is also used to convert slaked lime into lime.

Reactive

Don't decompose at all

Moderate

Form the metal oxide and water

Unreactive

Don't exist

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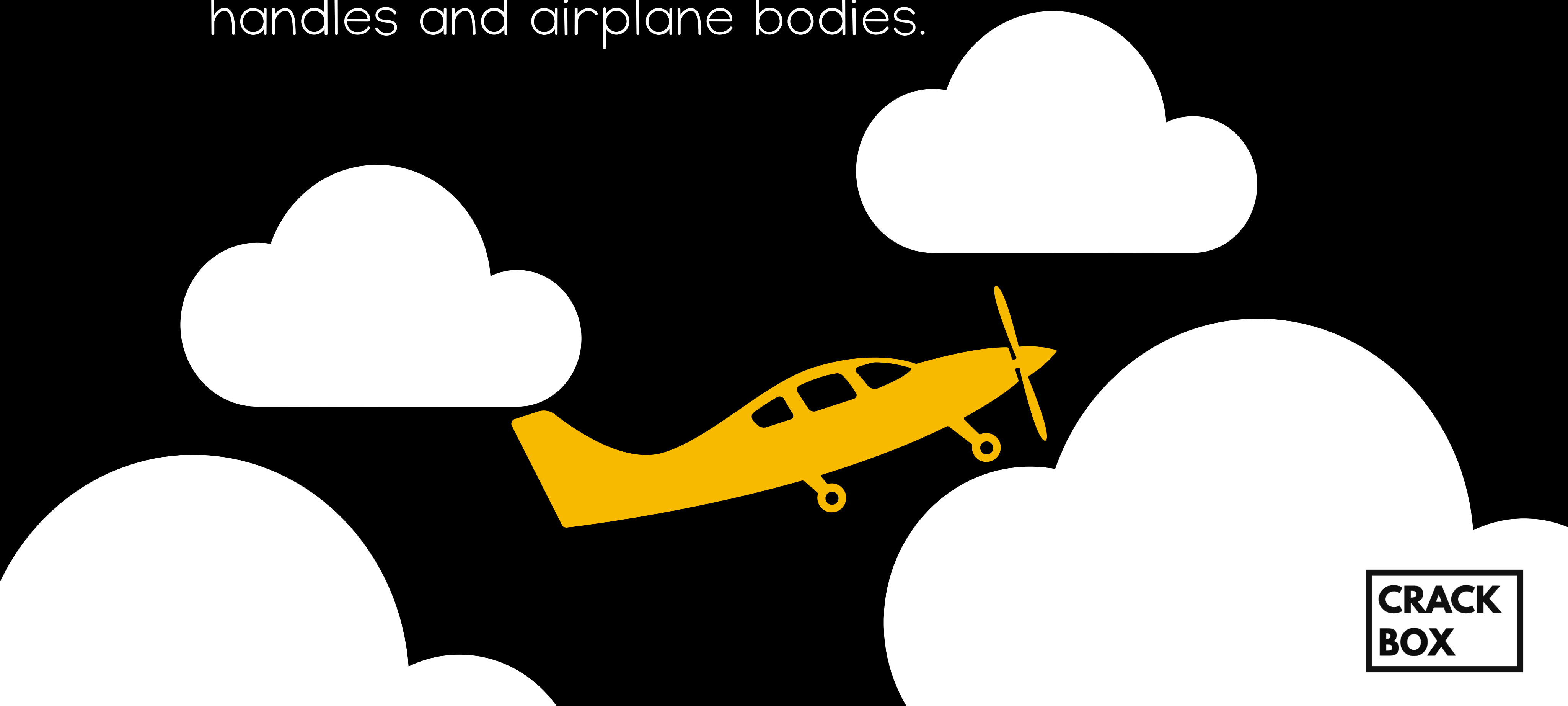
Competition Reactions

Generally, we find more uses for the unreactive metals like gold, etc. But aluminium is an exception.

Aluminium appears in the reactivity series right below magnesium and is quite reactive. Fortunately, it forms a thick oxide layer on its surface which prevents further reaction. This gives us the use of a light, strong metal for items such as door handles and airplane bodies.



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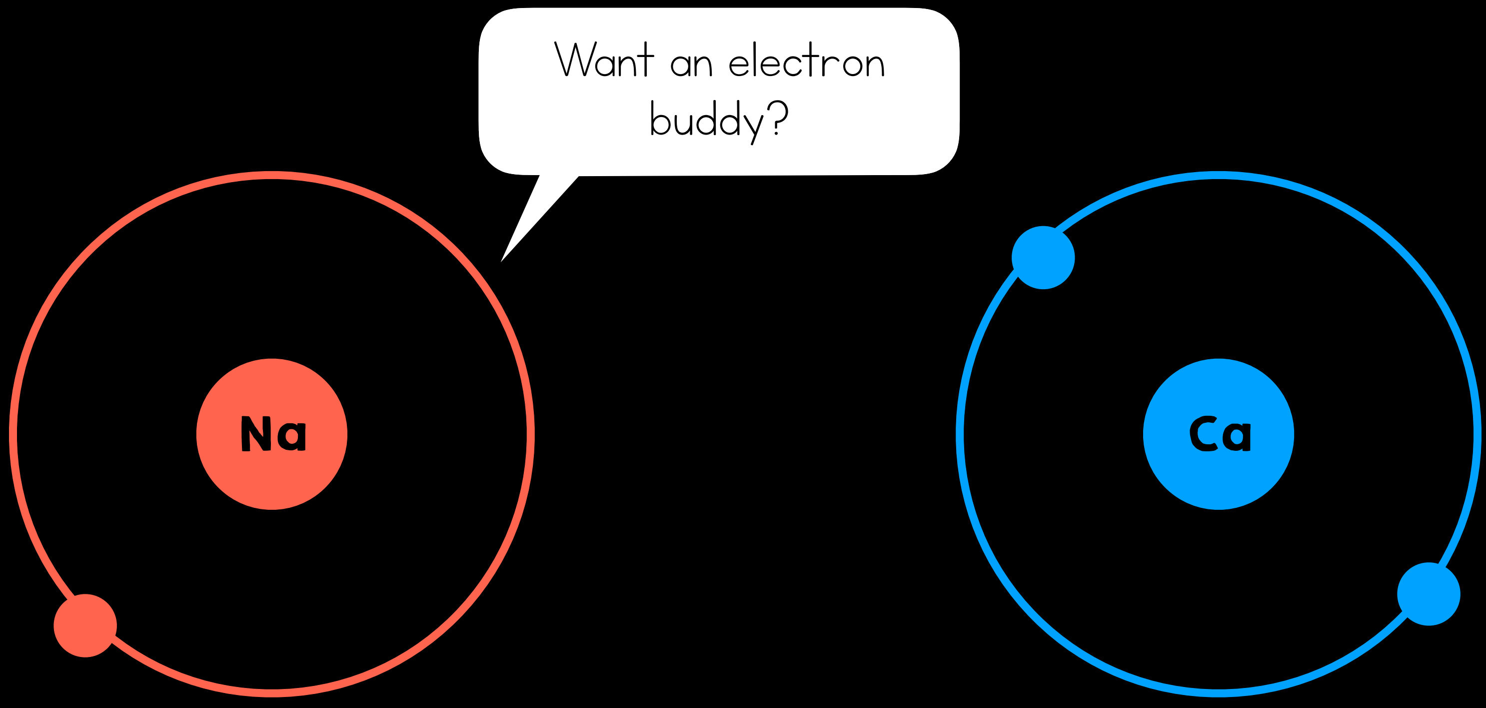
Competition reactions in the solid state

A more reactive metal has a greater tendency to form a metal ion by losing electrons than a less reactive metal.



Therefore, if a more reactive metal is heated with an oxide of a less reactive metal, then it will take the oxide from the less reactive metal.

For example lets take aluminium and iron(III) in a thermit reaction

The Aluminium being more reactive than Iron(III), takes the oxygen. This reaction leaves molten iron and a white powder of aluminium oxide. This reaction is also very exothermic making it ideal to weld damaged rail roads.

$$Fe_2O_3(s) + 2Al(s) \xrightarrow{\text{HEAT}} Al_2O_3(s) + 2Fe(s)$$

Competition reactions in the liquid state

Practically same like the thermit reaction, these reactions involve the metals competing for other anions. These reactions are called displacement reactions and the more reactive metal displaces the less reactive one from its salt.

For example lets take zinc and copper nitrate where zinc will displace copper to give zinc nitrate and copper.

$$Cu(NO_3)_2(aq) + Zn(s) \longrightarrow Zn(NO_3)_2(aq) + Cu(s)$$

Identifying Metal Ions

When an alkali hydroxide is added to a metal salt solution, a metal hydroxide is precipitated.

Now, based on the precipitate formed, we can predict which metal ion was present in the salt solution.

The following tables shows us the result of 7 metal ions when Sodium Hydroxide and Ammonium Hydroxide is added in a few drops, and later in excess.

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METAL ION PRESENT IN SOLUTION	Effect of Sodium Hydroxide Solution	
	Few Drops	In Excess
Aluminium	WHITE PPT	PPT DISSOLVES
Calcium	WHITE PPT	PPT DOESN'T DISSOLVE
Copper(II)	BLUE PPT	PPT DOESN'T DISSOLVE
Iron(II)	GREEN PPT	PPT DOESN'T DISSOLVE
Iron(III)	BROWN PPT	PPT DOESN'T DISSOLVE
Zinc	WHITE PPT	PPT DISSOLVES
Chromium	LIGHT GREEN PPT	PPT DISSOLVES

PPT = Precipitate

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METAL ION PRESENT IN SOLUTION	Effect of Aqueous Ammonia Solution	
	Few Drops	In Excess
Aluminium	WHITE PPT	PPT DOESN'T DISSOLVE
Calcium	NO PPT	NO PPT
Copper(II)	BLUE PPT	PPT DISSOLVES
Iron(II)	GREEN PPT	PPT DOESN'T DISSOLVE
Iron(III)	BROWN PPT	PPT DOESN'T DISSOLVE
Zinc	WHITE PPT	PPT DISSOLVES
Chromium	LIGHT GREEN PPT	PPT DISSOLVES (TO EXTENT)

PPT = Precipitate

Ammonia Solution = Ammonium Hydroxide

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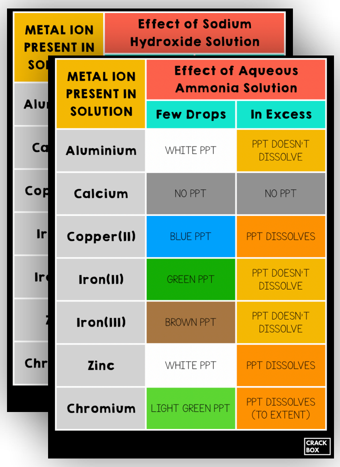

METAL ION PRESENT IN SOLUTION	Effect of Aqueous Ammonia Solution	
	Few Drops	In Excess
Aluminium	WHITE PPT	PPT DOESN'T DISSOLVE
Calcium	NO PPT	NO PPT
Copper(II)	BLUE PPT	PPT DISSOLVES
Iron(II)	GREEN PPT	PPT DOESN'T DISSOLVE
Iron(III)	BROWN PPT	PPT DOESN'T DISSOLVE
Zinc	WHITE PPT	PPT DISSOLVES
Chromium	LIGHT GREEN PPT	PPT DISSOLVES (TO EXTENT)

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The hydroxides of metals are basic and react with acids to form salts, however there are some metal hydroxides which also react with strong bases such as sodium hydroxide. Therefore called, amphoteric metals hydroxides.

Some metals who form amphoteric hydroxides

Pb

Zn

Al

The oxides of these metals are also amphoteric in nature and react with both, acids and alkalis

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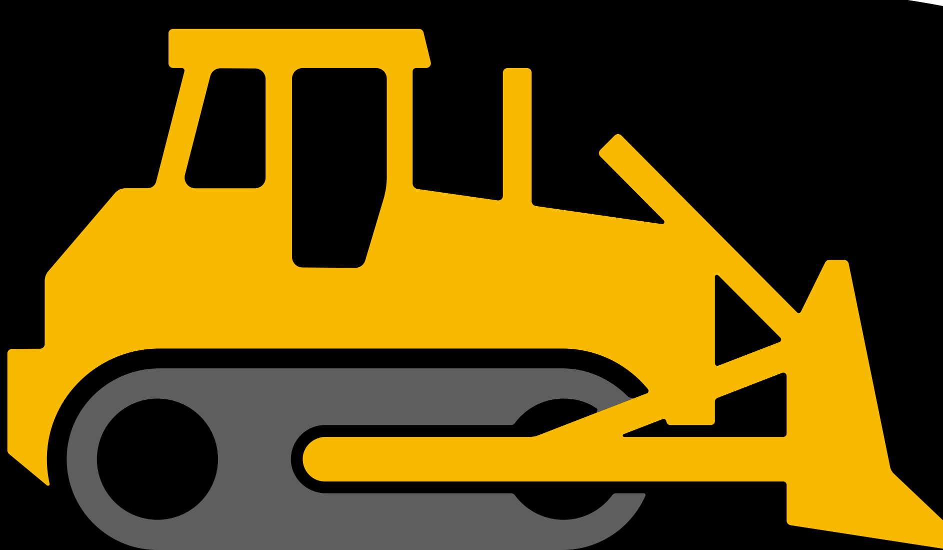
Extraction of Metals

Most metals are found in form of compounds called ores. These ores need to go through an extraction process to separate the metal from other substances and impurities.

The table on the next page shows us a brief about the metal and it's ore.



METAL	ORE	ORE COMPOUND	FORMULA	EXTRACTION METHOD
Aluminium	Bauxite	Aluminium Oxide	$Al_2O_3.2H_2O$	Electrolysis
Copper	Copper Pyrites	Copper Iron Sulfide	$CuFeS_2$	Roasting
Iron	Haematite	Iron (III) Oxide	Fe_2O_3	Heating with Carbon
Sodium	Rock Salt	Sodium Chloride	$NaCl$	Electrolysis
Zinc	Zinc Blende	Zinc Sulphide	ZnS	Roasting and Heating with Carbon

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Extraction of Iron

Iron is mainly found in its oxides, two examples being haematite and magnetite.

To separate iron from its oxides, the oxygen in the ore is made to react with carbon monoxide which acts like a reducing agent.

But it isn't that easy to get carbon monoxide, so a series of reactions need to be done in order to get carbon monoxide.

In an approximately 50m blast furnace, coke(C), limestone($CaCO_3$) and haematite(Fe_2O_3) are added with a temperature of upto 1100^0C .

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The reactions that happen inside

1. The coke and oxygen give carbon dioxide and limestone decomposes to give carbon dioxide and calcium oxide

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

$$CaCO_3\,(s) \longrightarrow CaO\,(s) + CO_2\,(s)$$

2. With all the carbon dioxide now, more coke reacts with the carbon dioxide and gives carbon monoxide

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

3. The carbon monoxide then does the job :)

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

Reducing Agent

Oxidising Agent

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Extraction of Zinc

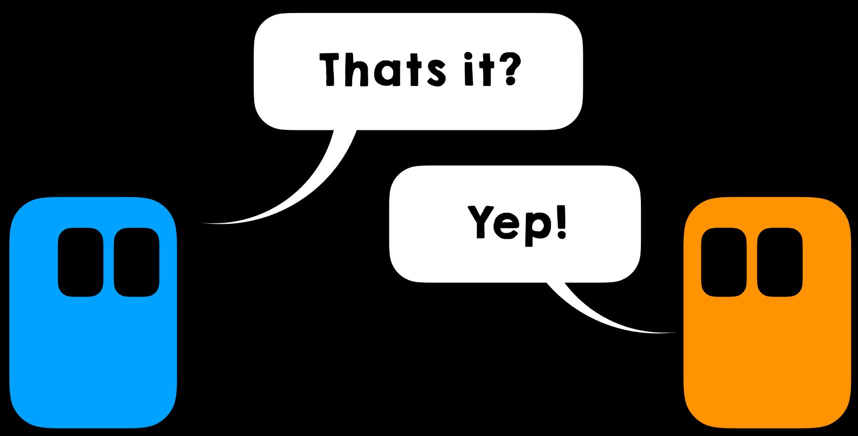
Zinc is much simpler than iron. It is mainly found as Zinc Sulphide or Zinc Blende (ZnS).

This ZnS is heated very strongly in an air current to convert it to Zinc Oxide (ZnO).

$$2ZnS(s) + 3O_2(g) \longrightarrow 2ZnO(s) + 2SO_2(g)$$

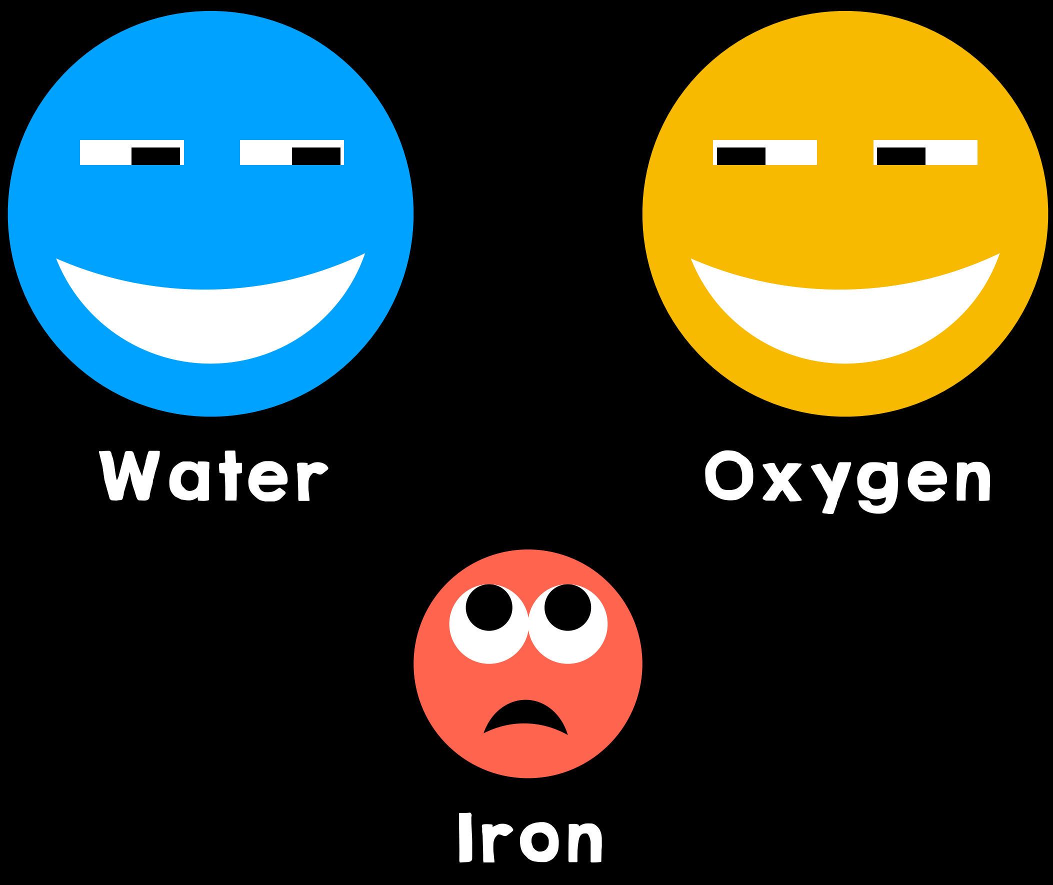
The oxide is then heated to about $1400^{\circ}C$ with coke.

$$ZnO(s) + C(g) \longrightarrow Zn(s) + CO(g)$$



RUSTING

After a period of time, when iron or steel is exposed to water and oxygen, a layer of rust develops, this orange-red powder causes expenditures of about one billion dollars world wide in replacing iron or steel structures! That's why it is important to prevent rusting and save costs.



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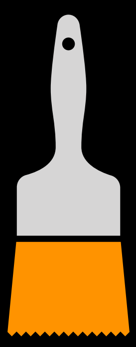
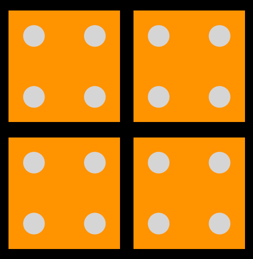
To prevent contact of oxygen and water with iron or steel, these techniques can be used to prevent maximum rusting

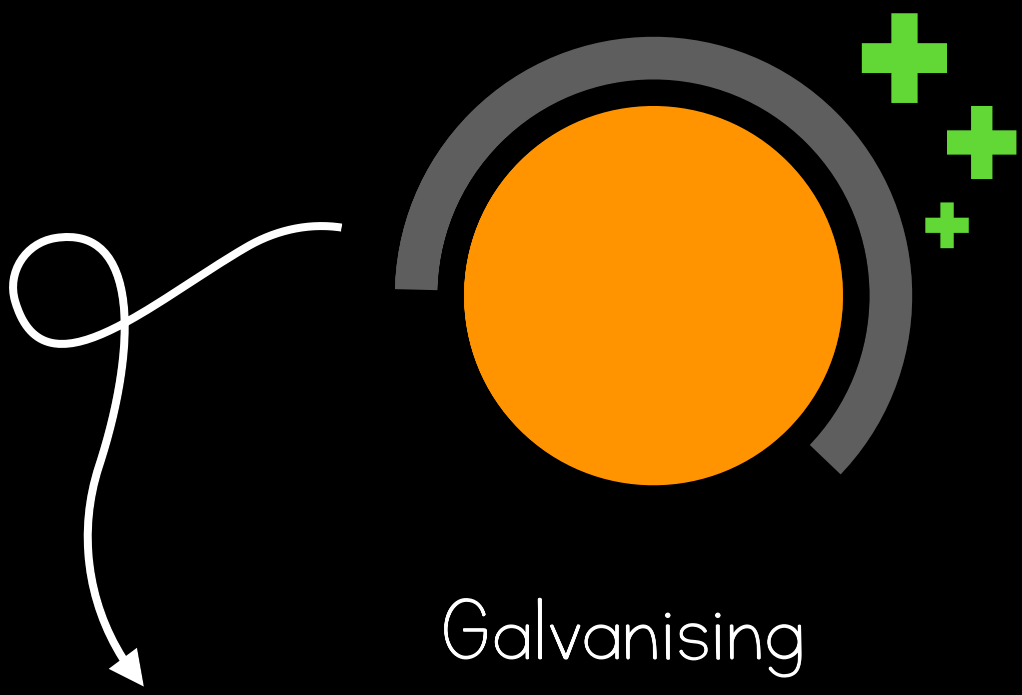
Painting



Greasing



Plating



Galvanising

Involves iron/steel to be dipped in <u>molten zinc</u> which gives it a protective layer as zinc is more reactive and corrodes itself but protects the iron/steel

ALLOYS

Alloys are <u>mixtures</u> of two or more metals (and sometimes include non-metals). They are usually formed by mixing molten metals thoroughly.

This usually results in substances that are much stronger and corrosion resistant than either of their parent metals, for example brass (copper and zinc).

ALLOY	COMPOSITION	USE
Brass	65% COPPER, 35% ZINC	Jewellery, machine bearings, electric connections, door furniture
Bronze	90% COPPER, 10% TIN	Casting, machine parts
Magnalium	70% ALUMINIUM, 30% MAGNESIUM	Aircraft construction
Pewter	30% LEAD, 70% TIN	Plates, drinking mugs, ornaments
Solder	70% LEAD, 30% TIN	Connecting electrical wires
Stainless Steel	74% IRON, 18% CHROMIUM, 8% NICKEL	Surgical instruments, cutlery, kitchen sinks

Not important to remember

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the end.