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The Reactivity Series

Metals are all not equally reactive. They can be arranged in order of their reactivity.

A metal 'higher' in the reactivity series:

- Reacts vigorously and quickly with chemicals to form products
- Corrodes more easily
- Loses electrons more readily to form positive ions
- Is more stable to the action of heat on its carbonates (e.g. K_2CO_3 and Na_2CO_3)
- Is extracted from its oxides using electrolysis (K to Al) as it is not possible with simple reduction methods using carbon and/or hydrogen

The reactivity series can be used in:

- Predicting what chemical reaction, if any, will happen
- Predicting stability of metal compounds
- The extraction of metals
- The prevention of rusting

Reaction of Metals with Water and Steam

- Metal + Water -> Metal Hydroxide + H_2 (g)
- Metal + Steam -> Metal Oxide + H_2 (g)

Metal	Reaction with	
	Cold Water	Steam
Potassium	Explosive Reaction	Explosive Reaction
	$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$	
Sodium	Very fast reaction, often explodes	Explosive Reaction
	$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$	
Calcium	Quick Reaction, plenty of gas bubbles produced	Explosive Reaction
	$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$	
Magnesium	Almost no reaction with cold water, a few gas bubbles produced very slowly	Vigorous reaction, heated magnesium burns, very bright flame
	$Mg(s) + 2H_2O(l) \downarrow Mg(OH)_2(aq) + H_2(g)$	$Mg(s) + H_2O(g) \downarrow MgO(s) + H_2(g)$
Zinc	No Reaction	Reacts readily, heated zinc burns, very bright flame
		$Zn(s) + H_2O(g) \downarrow ZnO(s) + H_2(g)$
Iron	No reaction with hot or cold water	Slow reaction (Reaction of red-hot iron)
		$3Fe(s) + 4H_2O(g) \downarrow Fe_3O_4(s) + 4H_2(g)$
Lead Copper Silver	No reaction under any conditions	No reaction under any conditions

Reaction of Metals with Dilute Acids

Only Metals above Hydrogen, H, in the reactivity series will react with dilute acids

Metal + Dilute Acid -> Salt + Hydrogen Gas

Metals below H in the reactivity series like Copper and Silver and Gold will not react with acids

Metal	Reaction with Hydrochloric Acid
Potassium Sodium	Explosive Reaction $2K(s) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2(g)$ $2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$
Calcium Magnesium	Very fast reaction, plenty of gas bubbles produced $Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$ $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
Zinc	Quick reaction, plenty of gas bubbles produced (less vigorous reaction compared to that with magnesium) $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
Iron	Slow reaction with cold acid, becomes quick when heated (produces gas bubbles and pale green solution) $Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$
Copper Silver	No reaction

Important Tips!

Lead reacts briefly with HCl and H_2SO_4 before the reaction stops.

Reason: Lead (II) Sulfate and Lead (II) Chloride forms an insoluble layer and prevents further reaction of acids with metal

Copper usually will not react with acids as copper is lower than hydrogen in the reactivity series. However:

$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$

Nitric acid acts as a strong oxidising agent. So this reaction is a redox reaction, rather than a acid-metal reaction

Aluminium, Al, does not react with water or steam as it is covered by a layer of aluminium oxide. This oxide is insoluble in water. This makes Al very passive. When Al is added to a dilute acid, the oxide layer slowly dissolves, exposing the metal. The metal then reacts with acid liberating hydrogen gas.

Displacement Reactions

A more reactive metal (higher in the reactivity series) will displace the less reactive metals (lower in the reactivity series) from their salts dissolved in water.

The displacement of metals from a solution can be used to find the relative position of metals in the reactivity series.

Metal	Observations recorded when metal is placed in aqueous solution of		
	Lead (II) Nitrate	Copper (II) Nitrate	Silver Nitrate
Iron	Grey deposit Light green solution formed	Reddish-brown deposit Blue Solution fades	Silvery Deposit Light Freen Solution formed
Lead	No visible change	Reddish-brown deposit Blue Solution fades	Silvery deposit
Copper	No visible change	No visible change	Silvery deposit Blue Solution formed

Important Recap!

Very often you will be asked to write ionic equation for displacement reactions!

e.g.

Thermal Stability of Compounds

<u>The more reactive the metal, the more stable the compound, the more difficult to decompose the compound.</u>

Sodium and metals that are above sodium in the reactivity series are generally very reactive and form stable carbonates which are not decomposed by heat.

Carbonates containing metals below sodium in the reactivity series can be decomposed by heat into metal oxide and carbon dioxide. In some cases, the metal oxide further decomposes to form the metal and oxygen gas if the metal is very low in the reactivity series

Metal carbonate -> Metal oxide + carbon dioxide

e.g. $CaCO_3$ (s) -> CaO (s) + CO_2 (g)

Decomposes more easily	Metal Carbonate	Effect of heat on the carbonate
↓	Potassium Carbonate Sodium Carbonate	No Change
	Calcium Carbonate Magnesium Carbonate Zinc Carbonate Iron (II) Carbonate Lead (II) Carbonate Copper (II) Carbonate	Decompose on heating, produces carbon dioxide and metal oxide $ZnCO_3$->ZnO+CO_2
	Silver Carbonate	$2Ag_2CO_3$ ↓ 4Ag + $2CO_2$ + O_2_

Reduction of Metal Oxides by Metal

A more reactive metal (higher up in the reactivity series) will reduce oxides of less reactive metals (lower in the reactivity series)

e.g. Zn (s) + CuO (s) -> ZnO (s) + Cu (s)

Extraction of Metals

- Metal ores are usually oxides, sulfides or carbonates mixed with other impurities
- Pure Metals must be extracted from their ores before they can be used in different industries
- There are 2 methods of extraction:
    - Electrolysis of molten ore
    - Reaction of metal ore with carbon (coke) / carbon monoxide or hydrogen
- The method of extraction depends on the position of the metal in the reactivity series

Reactivity decreases	Metal	Reactivity of metal	Method used to extract metal
↓	Potassium Sodium Calcium Magnesium Aluminium	Very reactive metals	Electrolysis (passing an electric current through molten compound)
	Zinc Iron Lead Copper Silver	Moderately reactive metals In the bottom of reactivity series	Reduction of metal oxides by hydrogen, carbon, or carbon monoxide Heating metal oxide with coke (carbon)

Reduction of Metal Oxides by Hydrogen, Carbon or Carbon Monoxide

The more reactive a metal is, the more stable is its oxide.

The oxides of highly reactive metals are not easily reduced to the metal.

Less reactive metals will form oxides which are relatively more unstable and can be reduced more readily

Metals below carbon e.g. Zinc, iron, tin, and lead in the reactivity series are extracted by the reduction method. The more reactive Carbon, C, is able to reduce oxides of these metals.

e.g. Lead is extracted by reducing Lead (II) Oxide using Coke.

2PbO (s) + C (s) -> 2Pb (s) + CO_2 (g)

Hydrogen, Carbon and Carbon Monoxide are reducing agents which are arranged in increasing reducing power.

Metal	Extraction Method
Potassium Sodium Calcium Magnesium Aluminium	These metals are very reactive and the metallic ions form strong ionic bonds with the oxide ions. They are not reduced by hydrogen, carbon and carbon monoxide. Electrolysis is used to obtain these metals from their molten oxides and salts instead.
Zinc	Zinc Oxide can be reduced to Zinc only by Carbon or Carbon Monoxide $ZnO + 2CO \rightarrow Zn + CO_2$ ZnO cannot be reduced by Hydrogen
Iron Lead Copper	These metal oxides can be reduced to their metals by hydrogen, carbon and carbon monoxide $CuO + H_2 \rightarrow Cu + H_2O$ $PbO + H_2 \rightarrow Pb + H_2O$ $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
Mercury Silver	Oxides of these metals are very unstable. They are easily decomposed by heat or light to give pure metals. No reducing agent is needed. $2HgO \rightarrow 2Hg + O_2$
Gold Platinum	These are very unreactive metals and found in its elemental state in nature

Rusting

Rusting is the slow oxidation of iron to form rust, a brown solid.

Rust is hydrated Iron (III) oxide $Fe_2O_3 . xH_2O$ where x indicated the number of water molecules surrounding each unit of iron (III) oxide.

Rusting Conditions: Oxygen and Water

Rust Prevention

- To stop or slow down rusting, water and air must be prevented from coming into contact with the surfaces of iron.
- Rusting can be prevented by 2 main methods: Barrier Method and Sacrifical Protection.

Barrier Method

Action	How it works	Examples of uses
Cover the iron surface with oil or grease	The layer of oil or grease prevents air and water from reaching the iron surface	Used in machines such as engines, bicycle chains
Painting iron surface	The layer of paint prevents air and water from reaching the iron.	Used for cars, iron railings, ships, bridges
Coating with another less reactive metal	Coating of metal prevents air and water from reaching the surface of the iron. However, if the surface is scratched , rusting will still occur	Used in cans in canned food industry

Sacrificial Protection

- Put the iron or steel in contact with a more reactive metal so that the more reactive metal corrodes instead of the iron or steel.
- Examples:
    - Blocks of Magnesium are attached to underground steel pipelines by an insulated electric cable. The more reactive magnesium corrodes instead of the steel pipelines.
    - Blocks of Zinc are placed on the hulls of ships. The more reactive Zinc corrodes instead of the iron.

Important Tip!

For Sacrificial Protection, you don't need to cover the entire metal with the sacrificial metal!
Physical Contact or via a wire is all that is needed!

Summary of Reaction of Metals with Cold Water, Steam and Acid

Metal	Metal with cold water	Metal with steam	Metal with acid
Potassium Sodium	Violent Reaction	Explosive Reaction	Explosive Reaction
Calcium	Rapid Reaction		Violent Reaction
Magnesium	Slow Reaction	Violent Reaction	Rapid Reaction
Aluminium Zinc	No Reaction	No reaction Rapid Reaction	Moderately Fast Reaction
Iron		Slow Reaction	Slow Reaction
Tin Lead		No Reaction	Very Slow Reaction
(Hydrogen) Copper Mercury Silver Gold Platinum			No Reaction

Summary of Reaction of Metal Oxides with Hydrogen and Carbon, and the action of heat on metal carbonates

Metal	Reduction of Metal Oxides with		Action of heat on metal carbonates
	Hydrogen	Carbon	
Potassium Sodium	Metal Oxides Cannot be reduced by Hydrogen	Metal Oxides cannot be reduced by Carbon	No decomposition
Calcium Magnesium Aluminium			Metal Carbonates are decomposed to form metal oxides and carbon dioxide Decomposition with increase ease down the series
Zinc		Metal Oxides reduced to form metal and carbon dioxide. Metal Oxides reduced with increase ease down the series	
Iron Lead Copper	Metal Oxides reduced to form metal and water Metal oxides reduced with increasing ease down the series		
Silver			Metal carbonates are decomposed to form metal, oxygen and carbon dioxide