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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 \rightarrow Metal Hydroxide + H_2 (g)
- Metal + Steam \rightarrow Metal Oxide + H_2 (g)

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

Reaction of Metals with Dilute Acids

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

Metal + Dilute Acid \rightarrow 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	Explosive Reaction
Sodium	$2K(s) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2(g)$ $2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$
Calcium	Very fast reaction, plenty of gas bubbles produced
Magnesium	$Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$ $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
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	No reaction
Silver	

Important Tips!

Lead reacts briefly with HCl and H₂SO₄ 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:



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

The more reactive the metal, the more stable the compound, the more difficult to decompose the compound.

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

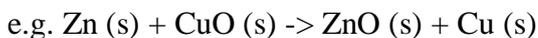
Metal carbonate \rightarrow Metal oxide + carbon dioxide

e.g. $\text{CaCO}_3 (\text{s}) \rightarrow \text{CaO} (\text{s}) + \text{CO}_2 (\text{g})$

Decomposes more easily ↓	Metal Carbonate	Effect of heat on the carbonate
	Potassium Carbonate	No Change
	Sodium Carbonate	
	Calcium Carbonate	Decompose on heating, produces carbon dioxide and metal oxide
	Magnesium Carbonate	
	Zinc Carbonate	$\text{ZnCO}_3 \rightarrow \text{ZnO} + \text{CO}_2$
	Iron (II) Carbonate	
	Lead (II) Carbonate	
	Copper (II) Carbonate	
	Silver Carbonate	$2\text{Ag}_2\text{CO}_3$ ↓ $4\text{Ag} + 2\text{CO}_2 + \text{O}_2$

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)



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:
 - o Electrolysis of molten ore
 - o 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	Very reactive metals	Electrolysis (passing an electric current through molten compound)
	Sodium		
	Calcium		
	Magnesium		
	Aluminium		
	Zinc	Moderately reactive metals	Reduction of metal oxides by hydrogen, carbon, or carbon monoxide
	Iron		
	Lead		
	Copper		
	Silver		
		In the bottom of reactivity series	Heating metal oxide with coke (carbon)

Reduction of Metal Oxides by Hydrogen, Carbon or Carbon Monoxide

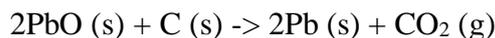
The more reactive a metal is, the more stable is its oxide.

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.



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

Metal	Extraction Method
Potassium Sodium	These metals are very reactive and the metallic ions form strong ionic bonds with the oxide ions.
Calcium Magnesium Aluminium	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 $Fe_2O_3 \cdot xH_2O$ 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 Sacrificial 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:
 - o Blocks of Magnesium are attached to underground steel pipelines by an insulated electric cable. The more reactive magnesium corrodes instead of the steel pipelines.
 - o 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	Violent Reaction	Explosive Reaction	Explosive Reaction
Sodium			
Calcium	Rapid Reaction		Violent Reaction
Magnesium	Slow Reaction	Violent Reaction	Rapid Reaction
Aluminium	No Reaction	No reaction	Moderately Fast Reaction
Zinc		Rapid Reaction	
Iron		Slow Reaction	Slow Reaction
Tin		No Reaction	Very Slow Reaction
Lead			
(Hydrogen)			No Reaction
Copper			
Mercury			
Silver			
Gold			
Platinum			

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