‘O’ Levels Pure Chemistry - Reactivity Series

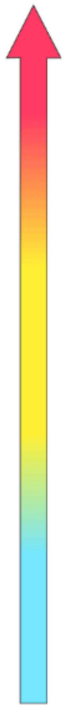
This article is written based on the 6092 Chemistry GCE Ordinary Level Syllabus.
https://www.seab.gov.sg/docs/default-source/national-examinations/syllabus/olevel/2024syllabus/6092_y24_sy.pdf

The reactivity series is a vital concept in O-Level Chemistry that ranks metals based on their reactivity with other substances, such as acids and water. Understanding the reactivity series is crucial for various applications, including metal extraction, corrosion prevention, and displacement reactions.

In this article, we will delve into the characteristics of the reactivity series, its significance, and how it influences chemical behavior in practical situations.

Metallic Reactivity Series

The metal reactivity series is a list of metals arranged in order of their reactivity, **starting with the most reactive** and **ending with the least reactive.** The position of a metal in this series determines how it reacts with water, acids, and other substances.

	Mnemonic	**Reactivity**	
Potassium (K)	Please	React with water	
Sodium (Na)	Stop		
Calcium (Ca)	Calling		
Magnesium (Mg)	Me	React with acid	
Aluminium (Al)	A		
(Carbon)	(Cool)		
Zinc (Zn)	Zebra		
Iron (Fe)	I		
Lead (Pb)	Like		
(Hydrogen)	(Happy)		
Copper (Cu)	Cute	Unreactive	
Silver (Ag)	Ant		

For O-Level Pure Chemistry, understanding the series helps you predict which metals can displace others in reactions and how they will interact with different chemicals.

A **more reactive metal** can displace a **less reactive metal** from its compound. For example, if you place a strip of zinc in a solution of copper(II) sulfate ($CuSO_4$), zinc will displace the copper:

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

Metal Reaction with Water, Steam and Acid

This series not only helps predict the outcomes of chemical reactions but also provides insights into the properties and behaviors of different metals.

Metals	Reaction with Cold Water	Reaction with Steam	Reaction with Acid	Thermal Decomposition of Carbonates
	Metal + cold water → metal hydroxide + hydrogen	Metal + steam → metal oxide + hydrogen	Metal + acid → salt + hydrogen	Metal carbonates → metal oxide + carbon dioxides OR metal + oxygen + carbon dioxides
Potassium	React violently to form metal hydroxide and hydrogen gas	React explosively to form metal oxide and hydrogen gas	React explosively to form salt and hydrogen gas	Does not decompose (stable to heat)
Sodium				
Calcium			React rapidly to form salt and hydrogen gas	Decompose to form metal oxide and carbon dioxide gas
Magnesium	React slowly (since magnesium hydroxide acts as protective barrier)	React violently to form magnesium oxide and hydrogen gas		
Aluminium	No reaction (since aluminium hydroxide acts as protective barrier)	No reaction (since aluminium oxide acts as protective barrier)		
(Carbon)				
Zinc	No reaction	React readily to form zinc oxide and hydrogen gas	React quickly to form zinc salt and hydrogen gas	Decompose to form metal oxide and carbon dioxide gas
Iron		React slowly to form iron (II/ III) oxide and hydrogen gas	Reacts slowly to form iron (II/III) salt and hydrogen gas	
Lead		No reaction	No reaction (since $PbCl_2$ is insoluble even though above H)	
(Hydrogen)				
Copper	No reaction	No reaction	No reaction	Decompose to form metal oxide and

				carbon dioxide gas
Silver				Decomponse to form metal, oxygen and carbon dioxide gas
Gold				
Platinum				

Metal Extraction

The extraction of metals from their ores varies depending on their position in the metal reactivity series. Metals higher in the series are more reactive and require more energy-intensive methods of extraction, while those lower in the series can be extracted more easily.

Metals	**Extraction Method**
Potassium	**Electrolysis**
Sodium	
Calcium	
Magnesium	
Aluminium	
(Carbon)	
Zinc	Reduction of metal oxides with **carbon**
Iron	
Lead	
(Hydrogen)	
Copper	Reduction of metal oxides with **hydrogen**
Silver	
Gold	Found in native state
Platinum	

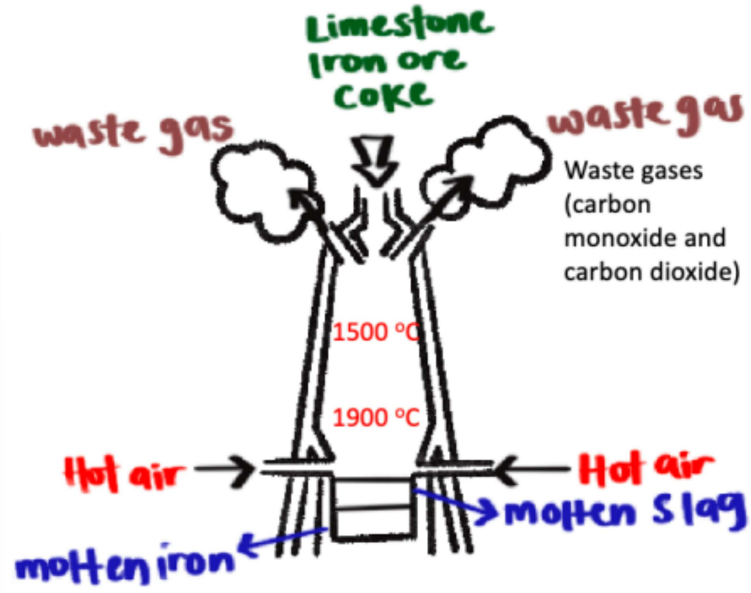
Let's look at a common example of metal extraction below.

Iron Extraction (Out of Syllabus)

Iron is one of the most widely used metals in the world. The extraction of iron using carbon is a process that has been practiced for centuries, forming the basis of steel production, which is vital for industries ranging from construction to manufacturing.

The extraction of iron using carbon occurs in a large furnace known as a **blast furnace**, which is designed to withstand extremely high temperatures. The key materials used in this process are:

- **Iron Ore (Fe_2O_3):** The source of iron.
- **Coke:** A form of carbon derived from coal, used as both a fuel and a reducing agent.
- **Limestone ($CaCO_3$):** Used to remove impurities from the ore by forming slag.
- **Hot Air:** Blown into the furnace to support combustion and provide the necessary heat.



Step 5: $CaO\ (s) + SiO_2\ (s) \rightarrow CaSiO_3\ (l)$
Molten slag flows to the bottom of the furnace remain on top of molten iron.

Step 4: $CaCO_3\ (s) \rightarrow CaO\ (s) + CO_2\ (g)$

Step 3: $Fe_2O_3\ (s) + 3CO\ (g) \rightarrow 2Fe\ (l) + 3CO_2\ (g)$
Molten iron flows to the bottom of the furnace.

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

Step 1: $C\ (s) + O_2\ (g) \rightarrow CO_2\ (g)$

Rusting

As we delve into the topic of iron, it's essential to also explore the phenomenon of rusting.

What is Rust?

Rusting is a common and destructive chemical process that occurs when iron is exposed to **moisture (H_2O)** and **air (O_2)** over time. This process leads to the formation of a reddish-brown substance known as **rust**, which can weaken and damage iron structures and objects.

$$\text{Iron} + \text{Oxygen} + \text{Water} \rightarrow \underset{\text{rust}}{\underline{\text{Hydrated iron (III) oxide}}}$$

$$4Fe(s) + 3O_2(g) + 2xH_2O(l) \rightarrow \underline{2Fe_2O_3.xH_2O(s)}$$

Methods to Prevent Rust

Given the damaging effects of rust, it is important to employ methods to prevent or slow down the rusting process. Here are some common strategies:

Methods	Description	Uses
Barrier (e.g., painting, coating with grease/oil)	Applying a layer of paint or oil provides a physical barrier that prevents moisture and oxygen from coming into direct contact with the iron	Cars, bridges, machines
Alloying (e.g., stainless steel)	Alloying iron with other metals enhances its resistance to rust. Chromium in stainless steel forms a thin, stable oxide layer on the surface that prevents further corrosion	Cutleries
Galvanising	Coating iron or steel with a layer of zinc, which acts as a sacrificial metal, meaning it corrodes in place of the iron	Buildings, signboards, sculptures
Sacrificial protection	Iron is connected to a more reactive metal, such as magnesium or zinc, which acts as a sacrificial anode. The more reactive metal corrodes instead of iron, thereby protecting it from rust.	Pipelines, water heaters, ship hulls

Final Comments

In conclusion, the reactivity series provides valuable insights into the behavior of metals in various chemical reactions and in practical applications, such as metal extraction and the prevention of corrosion. Mastery of the reactivity series will not only enhance your problem-solving skills, but also prepare you for advanced topics in chemistry. Keep exploring the fascinating world of chemistry, and let your curiosity guide you in discovering how these principles manifest in everyday life.

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