

NCERT Class 10 Science Notes

Chapter 3 – Metals and Non – Metals

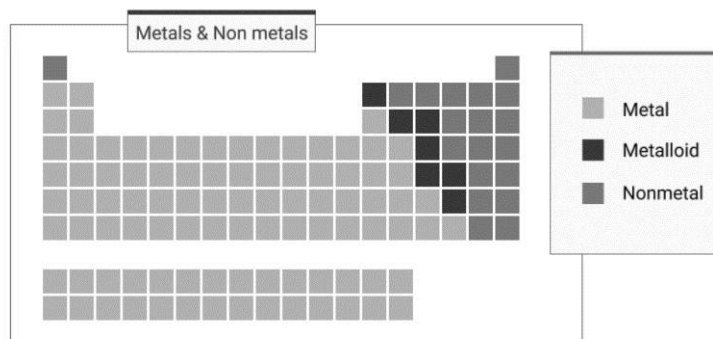
Introduction

The periodic table contains 118 elements, 92 of which are naturally occurring. The physical and chemical properties of metals and non-metals are vastly different. Currently, we are aware of roughly 80 metals.

Except for bromine, which is a liquid, over half of the non-metals are gases at ambient temperature. The most abundant non-metal in the earth's crust is oxygen, which makes up about half of the crust and is one of the main elements of air, along with nitrogen.

Silicon is the next most prevalent non – metal, accounting for around 26% of the earth's crust. Earth's two major elements are oxygen and silicon. The two principal ingredients of the oceans are hydrogen and oxygen.

Positions of Metals and Non – Metals in the Periodic Table



Metals are found in the periodic table's leftmost groupings. Group I A elements are alkali metals, which are highly reactive metals, and group II A elements are alkaline earth metals. Transition metals are elements that belong to groups II A and III A .

Non-metals are elements that are found to the right of the Periodic Table, in groups IV A , V A , VI A , and VII A , with the exception of hydrogen. From the top to the

bottom of the group, these elements get more non-metallic. The first and second members of group V A , for example, are non-metals, the third and fourth members are metalloids, and the last member is a metal. Metalloids are a class of elements that exhibit properties that are akin to both metals and non-metals. Boron, silicon, germanium, arsenic, antimony, tellurium, and astatine are the metalloids in question. The non-metals, which include the element hydrogen, are found to the right of these metalloids.

Group V A:

Non – Metals	Nitrogen, Phosphorous,
Metalloids	Arsenic, Antimony
Metal	Bismuth

Physical Properties of Metals

Physical State - With the exception of mercury and gallium, which are liquids at room temperature, metals are solids at room temperature.

Lustre – The property of metals which makes the light reflect from their surfaces is called lustre. This property of the metals can be due to the polished metal surfaces. Eg., gold and silver.

Malleability - Metals may be formed into thin sheets known as foils and can withstand hammering. With the exception of Zinc, which is fragile.

Ductility - Wires can be made out of metals. With the exception of Zinc, which is fragile.

Hardness - Except for sodium and potassium, which are soft and can be cut with a knife, all metals are hard.

Conduction - Because metals have free electrons, they are good conductors. Silver and copper are the best heat and electricity conductors. Lead is the least efficient heat conductor. Iron, bismuth, and mercury are likewise poor conductors.

Density - Metals have a high density and weigh a lot. The densities of iridium and osmium are the greatest, whereas lithium has the lowest density.

Melting and Boiling Point - Metals are known for their high melting and boiling points. The melting point of tungsten is the highest, while the boiling point of silver is the lowest. The melting values of sodium and potassium are both low.

Alloy Formation - Metals combine to create an alloy, which is a homogenous combination of metals. Brass is a copper and zinc alloy.

Sonorous – Metals, when hit by a solid object produces sound. This property of a metal is known as sonorous.

Physical Properties of Non – Metals

Physical State - At ambient temperature, the majority of non-metals exist in two of the three states of matter: gases (oxygen) and solids (iodine, carbon, sulphur). There is no metallic sheen to them (save iodine) and they do not reflect light. (With the exception of carbon in the form of diamond.)

Nature - Non – metals are extremely fragile, and they can't be coiled into wires or hammered into sheets. Except for diamond, which is the world's hardest substance.

Conduction - Non – metals are poor heat and electrical conductors. (Except graphite conducts heat, both graphite and gas carbon conduct electricity.)

Electronegative Character - Non – metals have a proclivity for gaining or sharing electrons with neighbouring atoms. Hence, non – metals are known for their electronegative nature.

Reactivity - When they come into contact with oxygen, they produce acidic or neutral oxides. Hence, non – metals are reactive.

Melting and Boiling Points - Non – metals are known for its low melting and boiling points.

Comparative properties of Metals and Non – metals

Property	Metals	Non – Metals
State of matter	Except for mercury, which is a liquid at ambient temperature, all are usually solid. Below 30°C , gallium and caesium dissolve. They may also be in a liquid form if the room temperature is approximately 30 C .	Non – metals exists in all three states. Only bromine is in the liquid state and iodine, carbon and sulphur are in solid states.
Density	Except sodium, potassium and calcium, rest all the metals have high density.	Non – metals have low densities.
Melting point	Metals are known to have high melting point except mercury, cesium, gallium, tin and lead.	The melting points of non – metals are low.
Boiling point	Metals are known to have high boiling points.	The boiling points of non – metals is low.
Hardness	All metals except mercury, sodium, calcium, potassium are hard.	Usually non – metals are not hard. The only exception is diamond, as it is the hardest substance on earth.
Malleability	Metals have the property that it can be beaten into sheets and hence they are malleable.	Non – metals cannot be beaten into sheets because they are brittle in nature.
Ductility	All the metals except sodium, potassium and calcium can be drawn into thin wires. Hence, they are ductile.	Non – metals cannot be drawn into thin wires. Hence, they are not ductile in nature.
Conduction of heat	Metals are good conductors of heat.	Non – metals are poor conductors of heat only except graphite.

Conduction of electricity	Metals are known to be good conductors of electricity.	Non – metals are non conductors of electricity only except graphite and gas carbon
Lustre	Metals which are cut freshly are lustrous in nature.	Non – metals usually do not have lustre only except diamond and iodine.
Alloy formation	Metals are known for their alloy formation.	Generally, non – metals do not form alloys. But carbon, phosphorous and sulphur are found in some alloys.
Tenacity	Metals usually have high tensile strength except sodium, potassium, calcium and lead.	Non – metals are known to have low tensile strength.
Brittleness	Although metals are hard but they are not brittle in nature, only except zinc which is brittle at room temperature.	Non – metals are generally brittle in nature.
Electronic configuration	Metals usually have the presence of 1, 2 or 3 electrons in their valence shell. The reactivity of the metals increases when the number of shells increases and the number of valence electrons are less.	Non – metals have the presence of 4, 5, 6, 7 electrons in their valence shells. If there are 8 electrons in the valence shell, then the element is called a noble gas. The reactivity of a non-metal increases as the number of shells decreases and the number of valence electrons increases.
Ionization	Metals are ionized when they lose electrons: $\text{Na} - e^- \rightarrow \text{Na}^+$	Non – metals are ionized when they gain electrons: $\text{Cl} + e^- \rightarrow \text{Cl}^-$
Charge on ions	Metals are always positively charge.	Non – metals are negatively charge species.

Type of valency	Metals are always electrovalent in nature.	Non – metals are both electrovalent and covalent.
Deposition during electrolysis	During electrolysis metals are always deposited at the cathode.	During electrolysis non – metals are always deposited at the anode.
Redox reaction	Metals tend to lose electrons and hence they get oxidized.	Non – metals tend to gain electrons and hence are reduced.
Redox agents	Metals are known as reducing agents.	Non – metals are known as oxidizing agents.
Nature of oxides	Metals usually form basic oxides and some of them are amphoteric in nature, such as: lead oxide, zinc oxide, aluminium oxide, etc.	Non – metals generally form acidic oxides. The neutral oxides formed by the non – metals are: nitrous oxide, carbon dioxide, nitric oxide, water, etc.
Hydrides	Except some transition metals, metals form hydrides.	Non – metals form hydrides, eg., NH_3 , PH_3 , HCl , HBr , HI , H_2S , H_2O
Atomicity	Metals are always mono-atomic.	Non – metals can be mono-, di-, tri-, or polyatomic.
Solubility	Metals tend to not dissolve in solvents. They are only dissolved by a chemical action.	Non – metals dissolve in the solvents and are re-obtained by the process of evaporation. Example: Sulphur in carbon disulphide.
Action with chlorine	Metals when reacted with chlorine, produced chlorides and these chlorides are electrovalent in nature.	Non – metals when reacted with chlorine also produce chlorides, but these chlorides obtained are covalent in nature.
Action with dilute acids	When metals are mixed with dilute acids, they give a respective salt and hydrogen.	Non – metals do not react with dilute acids.

Chemical Properties of Metals

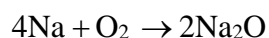
Metals are Electropositive Elements

Metals have a high reactivity. Metals are known as electropositive elements because they readily lose electrons and create positively charged ions. The metal sodium produces sodium ions (Na^+). Metals' electropositive nature makes it easy for them to form compounds with other elements.

Reaction of Metals with Oxygen

Some of the most reactive metals include sodium (Na) and potassium (K). In the presence of oxygen, potassium, sodium, lithium, calcium, and magnesium react and burn. Metals in the activity series of metals, from aluminium to copper, react slowly in air to generate metal oxides. Aluminium is the quickest, whereas copper is the slowest.

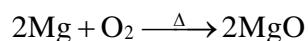
- At normal temperature, sodium metal interacts with oxygen in the air to generate sodium oxide. As a result, sodium is kept in kerosene to avoid reacting with oxygen, moisture, or carbon dioxide.



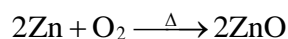
- Sodium oxide is a basic oxide that forms sodium hydroxide when it combines with water.



- At room temperature, magnesium does not react with oxygen. Magnesium burns in air with intense light and heat to create magnesium oxide (MgO) when heated.



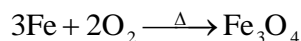
- Zinc when burned in air results in the formation of zinc oxide.



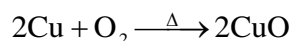
- Iron gets oxidized and forms rust in the presence of moist air.



- When iron is heated in the presence of air (oxygen) the formation of triferrous tetroxide takes place.



- Copper is the least reactive metal and, even when heated, does not burn in the air. Copper, on the other hand, reacts with oxygen and generates copper (II) oxide (CuO) on the surface and copper (I) oxide (Cu₂O) on the inside when heated for a long time.

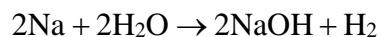


- When mixed with oxygen present in the air, gold and platinum do not react with it.

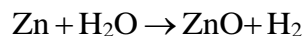
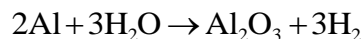
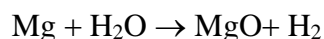
Reaction of Metal with Water

The metals that react with cold water are: potassium, sodium, lithium and calcium.

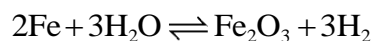
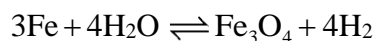
- Sodium when mixed with water is a highly vigorous reaction and forms sodium hydroxide and water.



- Metals in the activity series, from magnesium to iron, react with steam (but not cold water) to generate metal oxide and hydrogen gas.



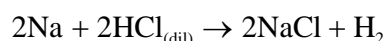
- When iron reacts with steam, it forms iron (I, II) oxide.



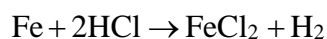
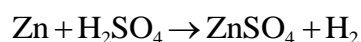
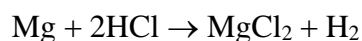
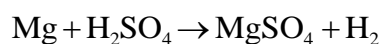
Note: The reaction between iron and steam is irreversible. Tin, lead, copper, silver, gold and platinum do not react with water or steam.

Reaction of Metals with Acids

- The metal salt (either sulphate or chloride) and hydrogen gas are formed when potassium, sodium, lithium, and calcium react forcefully with dilute H_2SO_4 and dilute HCl . The reaction is identical to the one that occurs when water is used.



- With dilute acid, magnesium, aluminium, zinc, iron, tin, and lead react safely. Magnesium is the fastest of the six elements, while lead is the slowest.



For the laboratory synthesis of hydrogen, zinc with dilute sulphuric acid is commonly utilised. At room temperature, the reaction is sluggish, but it can be sped up by adding a little copper (II) sulphate. Copper metal, which works as a catalyst, is displaced by zinc.

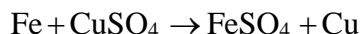
Copper, silver, gold, and platinum are metals that do not react with dilute acid to liberate hydrogen.

- Hydrochloric acid produces metal chlorides in general.
- Sulphuric acid reacts with metals to form sulphates.
- Nitric acid reactions are more complicated; nitrate is created, but the gas produced is rarely hydrogen and more frequently an oxide of nitrogen.

Reactions of Metals with Salt Solutions

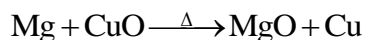
Reactive metals can displace any metal less reactive than themselves in solution or in their molten state, such as the less reactive metal's oxide, chloride, or sulphate. Metal A is more reactive than B if it displaces metal B from its solution.

Metal A + Salt solution of B → Salt solution of A + Metal B

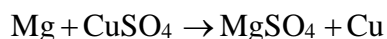







When dilute, copper (II) sulphate solution appears blue, while iron sulphate solution is practically colourless. The blue solution loses its colour during displacement, and the iron metal turns pink-brown as the displaced copper is deposited on it.

When magnesium powder and black copper (II) oxide are heated together, white magnesium oxide is created with brown copper bits:



The blue colour disappears as colourless magnesium sulphate forms and brown particles of copper metal form a precipitate when magnesium is added to a blue copper (II) sulphate solution:



Metal		Reaction with air	Reaction with water	Reaction with acids
Potassium	K	Burn vigorously to form oxides 	With cold water forms hydrogen gas and alkaline hydroxide solution. React with decreasing vigour down the series	Violent reaction to give hydrogen gas and salt solution
Sodium	Na			
Calcium	Ca	Burn with decreasing vigour down the series		
Magnesium	Mg			React to form hydrogen gas and salt solution with decreasing vigour down the series
Aluminium	Al		No reaction with cold water. With steam forms hydrogen gas and oxide	
Zinc	Zn			
Iron	Fe			
Lead	Pb			
Copper	Cu	React slowly on heating to form oxide layer		
Mercury	Hg		No reaction	No reaction
Silver	Ag			
Gold	Au	No reaction		

Electronic Nature of Metals and Non – Metals

With the exception of noble gases, all atoms have an incomplete outermost shell. Noble gases have a full outer shell and are hence non-reactive or "inert."

The stability of noble or inert gases is achieved via electron transfer or electron sharing in most reactive components. Metals are elements that can donate electrons. They lose electrons and create positive ions as a result.

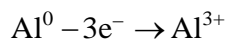
Non-metals are elements that can accept electrons. They gain electrons and generate negative ions as a result. The outermost shell of a metal has 1 to 3 electrons, while the outermost shell of a non-metal has 4 to 8 electrons.

Hydrogen and helium are the only two exceptions to this rule. Hydrogen is a non-metal with one electron in its valence shell, while helium has two electrons in its valence shell.

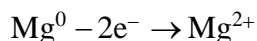
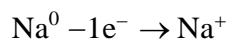
Types of Elements	Element	Atomic Number	Number of Electrons in the Shells				
			K	L	M	N	
Noble Gases	Helium (He)	2	2				
	Neon (Ne)	10	2	8			
	Argon (Ar)	18	2	8	8		
Metals	Sodium (Na)	11	2	8	1		
	Magnesium (Mg)	12	2	8	2		
	Aluminium (Al)	13	2	8	3		
	Potassium (K)	19	2	8	8	1	
	Calcium (Ca)	20	2	8	8	2	
	Non Metals	Nitrogen (N)	7	2	5		
		Oxygen (O)	8	2	6		
Fluorine (F)		9	2	7			
Phosphorous (P)		15	2	8	5		
Sulphur (S)		16	2	8	6		
	Chlorine (Cl)	17	2	8	7		

Important Points

Donor metals get a positive charge equal to the number of electrons they donate. Aluminum, for example, has an atomic number of 13, which means its electrical configuration is 2, 8, and 3. The valence shell of aluminium has three electrons; it loses three electrons to create Al^{3+} .

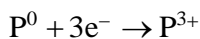
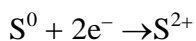
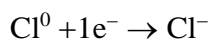


Other examples are given below:



Non – metals gain electrons and, as a result, gain a negative charge equal to the number of electrons accepted.

The examples are as follows:




The Reactivity Series of Metals

Despite the fact that most metals are electropositive in nature and lose electrons in chemical reactions, they do not respond with the same vigour or speed. Different metals react differently to different chemicals. The stronger an element's reactivity, the easier it is for it to lose its electrons and get a positive charge. Furthermore, the higher the number of shells and the lower the number of valence electrons, the higher the metal's reactivity. The metal activity series puts all metals in decreasing chemical activity order. The ease with which a metal loses electrons and generates positive ions in solutions reduces as we progress down the activity series from potassium to gold.

Potassium, the most active metal, is at the top of the list, while gold, the least reactive metal, is at the bottom. Despite the fact that hydrogen is a non-metal, it is included

in the activity series because it behaves like a metal in most chemical reactions, i.e., the hydrogen ion has a positive charge [H⁺] similar to other metals.

K	Potassium	
Na	Sodium	
Ca	Calcium	
Mg	Magnesium	
Al	Aluminium	
C	Carbon	
Zn	Zinc	
Fe	Iron	
Sn	Tin	
Pb	Lead	
H	Hydrogen	
Cu	Copper	
Ag	Silver	
Au	Gold	
Pt	Platinum	

Most reactive

Least reactive

C H added for comparison

Element	Symbol	Group Number
Potassium	K	I A
Sodium	Na	I A
Lithium	Li	I A
Calcium	Ca	II A
Magnesium	Mg	II A
Aluminium	Al	III A
Carbon	C	IV A
Zinc	Zn	II B
Iron	Fe	VIII
Tin	Sn	IV A
Lead	Pb	IV A

Hydrogen	H	I A
Copper	Cu	I B
Silver	Ag	I B
Gold	Au	I B
Platinum	Pt	VIII

- The more reactive a metal is in the series, the faster and more exothermic its reaction will be.
- This also means that the reverse reaction becomes more complex, implying that the more reactive a metal is, the more difficult it is to remove it from its ore. With oxygen and water, the metal is also more prone to corrosion.
- The reactivity series can be determined by watching metals react with water, oxygen, or acids.
- There are certain periodic table trends within the overall reactivity or activity series:

Metals	Reactivity and Reactions
Potassium	<p>Potassium is very reactive when mixed with cold water. The reaction yields, an alkali potassium hydroxide and hydrogen gas.</p> $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$
Sodium	<p>The reaction of sodium with cold water is very fast and the reaction forms sodium hydroxide and hydrogen gas.</p> $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$ <p>Sodium reacts with water by melting into a silvery ball that fizzes as it spins over the water. A colourless gas is produced by the quick exothermic reaction, which makes a squeaky pop! with a splint-hydrogen that is illuminated. The powerful alkali sodium hydroxide is created when the universal indicator changes from green to purple/violet. Because sodium is less thick than water, it floats.</p>

<p>Calcium</p>	<p>Calcium is quite reactive with cold water. On mixing calcium with cold water a moderately soluble alkali, calcium hydroxide and hydrogen gas is formed.</p> $\text{Ca (s)} + 2\text{H}_2\text{O (l)} \rightarrow \text{Ca(OH)}_2\text{(aq)} + \text{H}_2\text{(g)}$ <p>Calcium is very reactive with dilute hydrochloric acid. When mixed with dilute hydrochloric acid, it yields a colourless salt calcium chloride and hydrogen gas.</p> $\text{Ca (s)} + 2\text{HCl (l)} \rightarrow \text{CaCl}_2\text{(aq)} + \text{H}_2\text{(g)}$ <p>Calcium is not very reactive with dilute sulphuric acid as when the two reactants are mixed the calcium sulphate formed as a result is not very soluble and it coats the metal inhibiting the reaction.</p> $\text{Ca (s)} + \text{H}_2\text{SO}_4\text{(l)} \rightarrow \text{CaSO}_4\text{(s)} + \text{H}_2\text{(g)}$
<p>Magnesium</p>	<p>Magnesium reacts slowly with water to generate magnesium oxide, a weak soluble base, and hydrogen gas. The reaction with heated magnesium is faster with steam, and a white powder magnesium oxide is generated together with hydrogen. If magnesium has been previously ignited in air, it will burn with a dazzling white flame in steam.</p> $\text{Mg (s)} + \text{H}_2\text{O (l)} \rightarrow \text{MgO (s)} + \text{H}_2\text{(g)}$ <p>Magnesium burns in carbon dioxide and forms black specs of carbon.</p> $2\text{Mg (s)} + \text{CO}_2\text{(g)} \rightarrow 2\text{MgO (s)} + \text{C (s)}$ <p>Magnesium is very reactive with dilute hydrochloric acid and forms a colourless soluble salt, magnesium chloride and hydrogen gas.</p> $\text{Mg (s)} + 2\text{HCl (aq)} \rightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{(g)}$ <p>Magnesium is very reactive with dilute sulphuric acid and forms a colourless soluble salt, magnesium sulphate with hydrogen gas.</p>

	$\text{Mg (s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{MgSO}_4(\text{aq}) + \text{H}_2(\text{g})$
Aluminium	<p>Due to a protective aluminium oxide layer of Al_2O_3, aluminium has no interaction with water or steam. Aluminum chloride, a colourless soluble salt, and hydrogen gas are formed after a slow reaction with weak hydrochloric acid.</p> $2\text{Al (s)} + 6\text{HCl (aq)} \rightarrow 2\text{AlCl}_3(\text{aq}) + 3\text{H}_2(\text{g})$ <p>The reaction of aluminium with dilute sulphuric acid is an extremely slow reaction which forms a colourless salt aluminium sulphate and hydrogen gas.</p> $2\text{Al (s)} + 3\text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Al}_2(\text{SO}_4)_3(\text{aq}) + 3\text{H}_2(\text{g})$
Carbon (non – metal)	<p>Electrolysis (or substituting it with an even more reactive metal) is required to extract elements higher than carbon, such as aluminium and the more reactive metals. By reducing the heated metal oxide with carbon, metals underneath it, such as zinc or a less reactive metal, can be removed.</p>
Zinc	<p>Zinc does not react with cold water but when the reaction mixture is heated, the steam from the water converts zinc to zinc oxide and hydrogen gas is produced.</p> $\text{Zn (s)} + \text{H}_2\text{O (g)} \rightarrow \text{ZnO (s)} + \text{H}_2(\text{g})$ <p>Zinc is highly reactive with dilute hydrochloric acid and forms a colourless soluble salt zinc chloride with hydrogen gas.</p> $\text{Zn (s)} + 2\text{HCl (aq)} \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2(\text{g})$ <p>Zinc when reacts with dilute sulphuric acid and forms a colourless soluble salt, zinc sulphate with hydrogen gas.</p> $\text{Zn (s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{ZnSO}_4(\text{s}) + \text{H}_2(\text{g})$ <p>The above reaction can be catalyzed by adding a very small quantity of copper sulphate in the reaction mixture.</p>

	<p>The extraction process of zinc can be carried out by reducing the hot zinc oxide on heating with carbon.</p> $2\text{ZnO (s)} + \text{C (s)} \rightarrow 2\text{Zn (s)} + \text{CO}_2(\text{g})$
Iron	<p>Electroplating is also used to prevent corrosion by coating the surface with metals such as tin, chromium, and nickel. Steel cans are protected by a thin layer of un – reacted tin that lasts as long as the tin layer is complete.</p> $3\text{Fe (s)} + 4\text{H}_2\text{O (l)} \rightarrow \text{Fe}_3\text{O}_4(\text{s}) + 4\text{H}_2(\text{g})$ <p>Iron when mixed with dilute hydrochloric acid reacts slowly and a soluble pale green salt Iron (II) chloride salt along with hydrogen gas is formed.</p> $\text{Fe (s)} + 2\text{HCl (aq)} \rightarrow \text{FeCl}_2(\text{aq}) + \text{H}_2(\text{g})$ <p>When dilute acid is added to iron, it forms a soluble pale green salt Iron (II) sulphate and hydrogen gas.</p> $\text{Fe (s)} + 2\text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{FeSO}_4(\text{aq}) + \text{H}_2(\text{g})$ <p>The heated metal oxide can be reduced with carbon monoxide generated from carbon in the blast furnace to remove iron. Example:</p> $\text{Fe}_2\text{O}_3(\text{s}) + 3\text{CO (g)} \rightarrow 2\text{Fe (s)} + 3\text{CO}_2(\text{g})$ $\text{Fe}_3\text{O}_4(\text{s}) + 4\text{CO (g)} \rightarrow 3\text{Fe (s)} + 4\text{CO}_2(\text{g})$
Hydrogen (non – metal)	<p>None of the metals below hydrogen can create hydrogen gas when they react with acids. They are the least corroded metals, which contributes to their high value and use in jewellery, electrical connections, and other applications.</p>
Copper	<p>Copper has no reaction when it comes into contact with cold water or when it is heated in steam. There is no reaction with dilute hydrochloric or sulphuric acid. Copper may be extracted by heating the heated black metal oxide with charcoal and reducing it.</p>

	$2\text{CuO (s)} + \text{C (s)} \rightarrow 2\text{Cu (s)} + \text{CO}_2\text{(g)}$
Silver	There is no reaction with cold water or steam. There is no reaction with dilute hydrochloric or sulphuric acid. Silver can be extracted through reduction, but it can also be found as a 'native' element.
Gold	There is no reaction with cold water or steam. There is no reaction with dilute hydrochloric or sulphuric acid. Gold can be easily recovered from its ores via reduction, but it is usually found in its natural state. Gold is in its purest form when it is 24 carat.
Platinum	Platinum does not react with cold water or steam. There is no reaction with dilute hydrochloric or sulphuric acid. It is a relatively rare metal, similar to gold. It's found in high-end jewellery, industrial catalysts (inert crucible containers), and catalytic converters in automobile exhaust.

Bonding

The quantity of valence electrons determines an atom's willingness to participate in chemical reactions (electrons in the outermost shell of an atom). During chemical combination, atoms acquire the stable noble gas configuration of eight electrons in the outermost shell (known as the octet rule).

A combination of atoms can take place in one of two ways: electrovalent bonding or covalent bonding. In all chemical reactions, the electrons from an atom's outermost shell are involved in interacting with other atoms, either via transfer or sharing.

Electrovalent Bonding

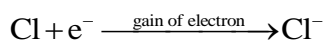
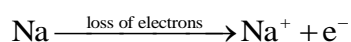
Electrovalency occurs when one atom donates one, two, or three electrons from its valence shell to another atom that has the ability to accept these electrons. Both of these atoms have the structure of an inert gas as a result of electrovalency. It is called ionic or electrovalent bond when a chemical link is formed by the transfer of electrons from one element's atom to the atom or atoms of another.

In NaCl, sodium has an electrovalency of 1+ and chlorine has an electrovalency of 1-. Calcium and magnesium, in their chloride form, have an electrovalency of 2+.

There are numerous elements with varying electrovalencies in various compounds. This phenomenon is known as 'varying electrovalency,' for example, iron resides in ferrous sulphate and ferric sulphate as Fe^{2+} and Fe^{3+} , respectively.

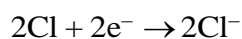
Formation of Sodium Chloride

Because sodium has only one electron in its valence shell, it loses one electron to complete its octet during the creation of an ionic bond between the metal sodium and the non-metal chlorine. It takes on the form of a noble gas, neon (2, 8). The chlorine atom contains seven electrons in its valence shell and gains one to complete its octet, as well as acquiring stable electronic configurations.

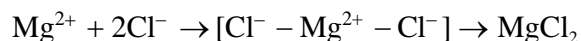


Formation of Magnesium Chloride

Magnesium, with an atomic number of 12, has electronic configuration as 2, 8, 2. It has two electrons in its valence shell. Chlorine's electronic configuration (atomic number: 17) is 2, 8, 7. There are seven valence electrons in it. Because magnesium has two electrons more than the neon configuration (2, 8) and chlorine has one electron less than the argon configuration (2, 8, 8), one atom of magnesium will seek out two atoms of chlorine to transfer its two electrons to (one to each), as seen below:



The ion of magnesium (Mg^{2+}) and the two ions of chlorine (Cl^-) form magnesium chloride which has an ionic bond between them.

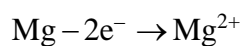


In the terms of Lewis dot structure

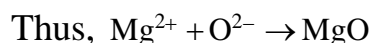
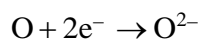


Formation of Magnesium Oxide

The atomic number of magnesium is 12 having an electronic configuration of (2, 8, 2). Magnesium atom loses two atoms to become stable like neon (2, 8)

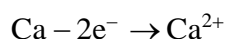


Oxygen having an atomic number 8 has the electronic configuration (2, 6). The oxygen atom gains two atoms so that it can become stable.

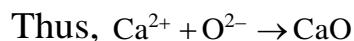
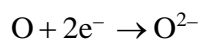


Formation of Calcium Oxide

The atomic number of calcium is 20 and its electronic configuration is (2, 8, 8, 2). To become stable calcium loses two electron, to achieve an electronic configuration of (2, 8, 8).



Oxygen having an atomic number 8 has the electronic configuration (2, 6). The oxygen atom gains two atoms so that it can become stable.



Properties of Electrovalent Compounds

Property	Electrovalent or Ionic Compounds
Structure of charge ions	Electrovalent compounds consist of oppositely charged molecules.
Physical state and hardness	Electrovalent compounds have a high inter-atomic attraction, making them brittle, rigid, crystalline solids.

Melting and boiling points	Because of the strong affinity between the particles of electrovalent compounds, melting or boiling them necessitates high temperatures.
Solubility	Electrovalent compounds are usually soluble in water, but in organic solvents they are insoluble.
Passage of electricity	Because ions cannot move in the solid state due to their rigid structure, ionic compounds do not carry electricity in the solid state. Because the electrostatic forces of attraction between oppositely charged ions are overcome in molten form or aqueous solution form, they enable the flow of electricity and are dissolved by it.
Rate of reaction	The rate of reaction of electrovalent compounds is very high that means that the reaction occurs at very high speed.
Dissociation in solution	Electrovalent substances dissociate in solution to produce negative and positive ions because they are made up of charged ions. $\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^-$
Electrolysis	Electrolysis is possible with electrovalent substances. At the cathode, cations are discharged, while anions are discharged at the anode.

Covalent Bond

The force of attraction that results from the mutual sharing of electrons between two nonmetallic atoms is known as a covalent bond. One, two, or three pairs of electrons may be shared by the combining atoms. A covalent bond is created between two nonmetal atoms that are similar or distinct by the mutual sharing of electrons, which contribute to the stability of both atoms.

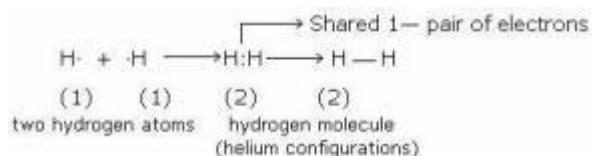
When two atoms unite through mutual electron sharing, each atom adopts the stable configuration of the neighbouring noble gas. Covalent compounds are those that form as a result of covalent bonding. Bond Pairs are the common pair of electrons.

Formation of Covalent Bonds

The Hydrogen Molecule

In the K shell, the hydrogen atom (atomic number=1) contains one electron. It tries to obtain helium's configuration (Atomic number 2). This is conceivable if the valence electrons of the two joining atoms are shared to form a single covalent bond.

The hydrogen molecule has one covalent bond between two hydrogen atoms and forms H_2 .



The Oxygen Molecule

The valence electrons in oxygen (atomic number 8) are six, two short of the octet configuration. Two pairs of electrons are shared by the two oxygen atoms, resulting in two covalent connections.

Covalency

The covalency of a compound is the number of electrons that an atom provides to mutual sharing during the creation of a chemical bond. Thus, hydrogen has a covalency of one in H_2 , oxygen has a covalency of two in O_2 , and nitrogen has a covalency of three in N_2 .

Lone pairs, also known as non-bonding pairs of electrons, occur when one or more pairs of electrons in the valence shell of an atom do not participate in bonding.

Multiple Covalent Bond

Multiple covalent bonds are covalent bonds formed by the mutual sharing of more than one pair of electrons. These are the following:

Double Covalent Bond

A double covalent bond, or simply a double bond, is the link formed between two atoms as a result of the sharing of two electron pairs. Two small horizontal lines (=) are put between the two atoms to denote it. Example: $O=O$, $O=C=O$ etc.

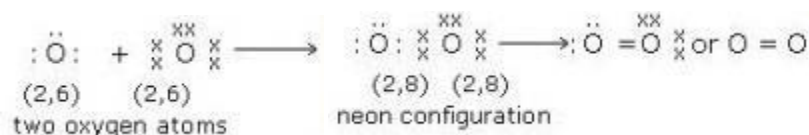
Triple Covalent Bond

A triple covalent bond, or simply a triple bond, is created when three electron pairs are shared. A triple bond is defined by three short horizontal lines connecting the two atoms. Example: $\text{N} \equiv \text{N}$ and $\text{H}-\text{C} \equiv \text{C}-\text{H}$

Formation of Molecules Having Double Bond

Formation of Oxygen (O_2) Molecule

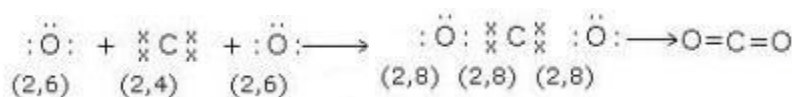
The valence shell of an oxygen atom has six electrons. As a result, two more electrons are required to obtain the closest noble gas arrangement. This is accomplished by the two oxygen atoms sharing two pairs of electrons, as seen below:



Formation of Carbon Dioxide Molecule (CO_2)

The electronic configuration of carbon (Atomic number 6) is (2, 4) and the electronic configuration of oxygen (Atomic number 8) is (2, 6).

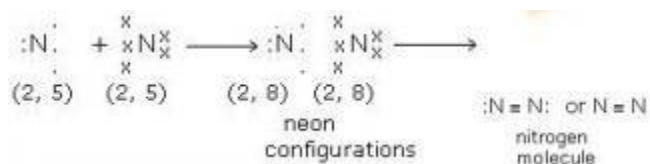
To get noble gas configurations, each carbon atom requires four additional electrons, and each oxygen atom requires two additional electrons. This is accomplished in the following way:



Formation of Molecules having Triple Bond

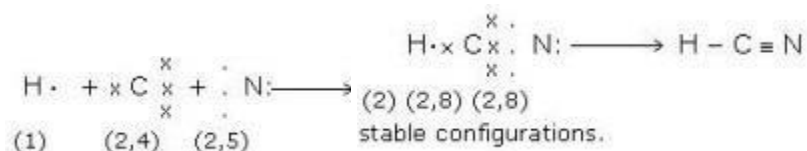
Formation of Nitrogen (N_2) Molecule

The valence shell of a nitrogen atom has five electrons. As a result, three additional electrons are required to obtain a stable configuration of the nearest noble gas (neon). This is accomplished by sharing three pairs of electrons in the manner described below:



Formation of Hydrogen Cyanide (HCN) Molecule

In HCN, the carbon atom shares one electron pair with hydrogen, resulting in a single covalent connection with the hydrogen atom. To form a triple bond between carbon and nitrogen, the carbon atom shares three electron pairs with the nitrogen atom. The atoms and Lewis structure of the HCN molecule are as follows:



General Properties of Covalent Compounds

The main properties of covalent compounds are:

- **State of existence:** Covalent chemicals exist as molecules rather than as ions. Because the molecules have weak intermolecular interactions, they exist as liquids or gases at normal temperature. A few chemicals, such as urea and sugar, do exist in the solid state.
- **Low melting and boiling points:** Covalent compounds often have low melting and boiling points. This is due to the weak forces that exist between the molecules, which are easily overcome at low temperatures.
- **Solubility:** In water and other polar solvents, covalent molecules are usually insoluble or just slightly soluble. Non-polar solvents like benzene and carbon tetrachloride, on the other hand, are soluble in them.
- **Non – conductors:** Because covalent compounds do not produce ions in solution, they are poor electrical conductors when fused or dissolved.
- **Molecular reactions:** Covalent compound reactions take place between their molecules. These include the breaking of covalent bonds in reacting molecules and the formation of new covalent bonds in product molecules.

Because energy is required to break covalent bonds, these processes are relatively sluggish.

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Occurrence of Metals

Minerals and Ores

Minerals are metals and their compounds that are present in the earth's crust. Ores are minerals from which metals can be profitably and easily mined. Metal complexes with a lower amount of impurities are found in ore. Therefore, it is clear that all ores are minerals, but not all minerals are ores.

Oxides	Carbonates	Halides	Sulphides	Sulphates
Zincates (ZnO)	Marble or limestone ($CaCO_3$)	Fluorspar (CaF_2)	Zinc Blend (ZnS) Galena (PbS)	Anglesite ($PbSO_4$)
Haematite ($Fe_2O_3 \cdot xH_2O$) Magnetite (Fe_3O_4)	Calamine ($ZnCO_3$)	Cryolite (Na_3AlF_6)	Iron Pyrites (FeS_2)	Barium Sulphate ($BaSO_4$)
Bauxite ($Al_2O_3 \cdot xH_2O$)	Siderite ($FeCO_3$)	Horn Silver ($AgCl$)	Cinnabar (HgS)	Gypsum ($CaSO_4 \cdot 2H_2O$)
Cuprite (Cu_2O)	Magnesite ($MgCO_3$)	Rock Salt ($NaCl$)		Epson Salt ($MgSO_4 \cdot 7H_2O$)

In the Free State

Only a few metals exist in their natural condition. Only pure metals such as gold, platinum, and mercury are occasionally found in the Free State. Copper and silver are occasionally discovered in the Free State. Air and water have no effect on such metals.

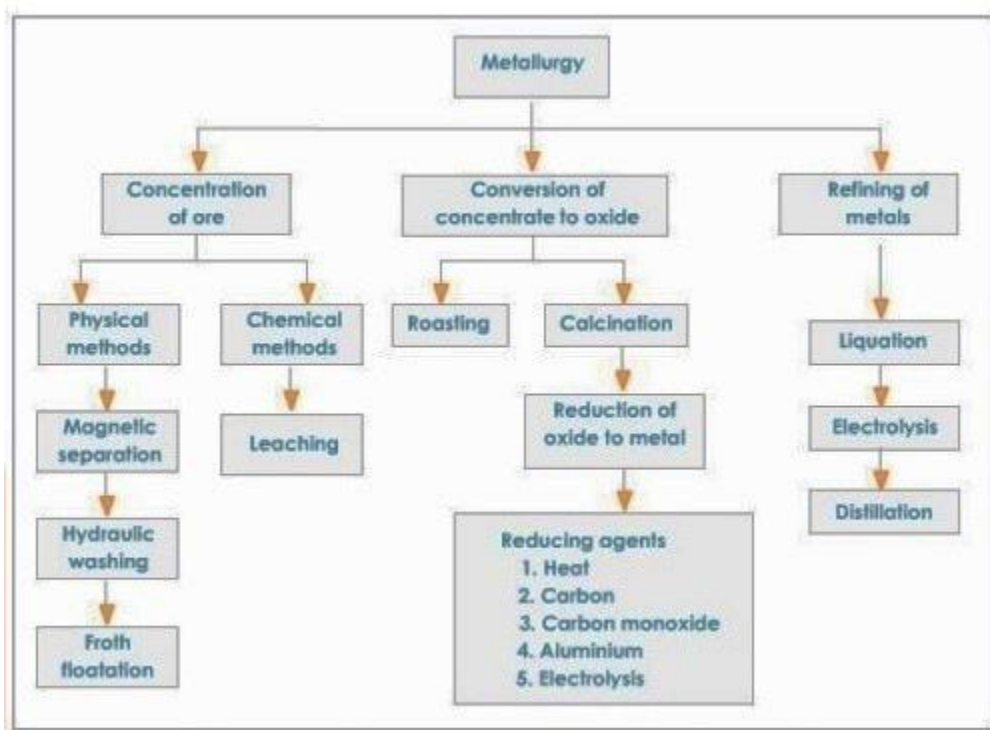
In Combined State

The remaining metals are found as oxides, carbonates, sulphides, sulphates, silicates, chlorides, nitrates, phosphates, and other compounds. Copper and silver are two metals that may be found in both free and mixed forms in the form of sulphide, oxide, or halide ores. Because the metals at the top of the activity series (K, Na, Ca, Mg, Al) are so reactive, they are never found as free elements in nature.

The intermediate metals in the activity series (Zn, Fe, Pb) are moderately reactive. They are mainly found as oxides, sulphides, or carbonates in the earth's crust.

Extraction of Metals – Metallurgy

Metallurgy is the study of the different processes involved in extracting metals from their ores and purifying them.



Concentration – Enrichment of Ores

Ore is an impure metal with a lot of sand and stony particles in it. The impurities in the ore, such as sand, stony minerals, limestone, mica, and so on, are referred to as gangue or matrix. Before the metal can be extracted, these contaminants must be

eliminated from the ore. A fusible compound called slag is formed when a chemical called flux is introduced to the ore to remove the matrix.

Gangue + Flux = Slag

The procedures for separating the gangue from the ore are based on the physical or chemical differences between the gangue and the ore. The ore is first crushed to a powder. Physical techniques such as hydraulic washing, froth flotation, and magnetic separation, as well as chemical processes, are used to separate the crushed ore, depending on the type of the ore and its impurities. Ore concentration is sometimes known as ore 'dressing' or 'enrichment.'

Physical Methods of Concentration

Hydraulic Washing (Gravity Separation)

The ore particles are put through a hydraulic classifier, which is a vibrating sloping table with grooves that is subjected to a stream of water. The heavier gangue particles are washed away, while the denser ore settles in the grooves. Heavy oxide ores of lead, tin, iron, and other metals are concentrated using this technique.



Froth Floatation

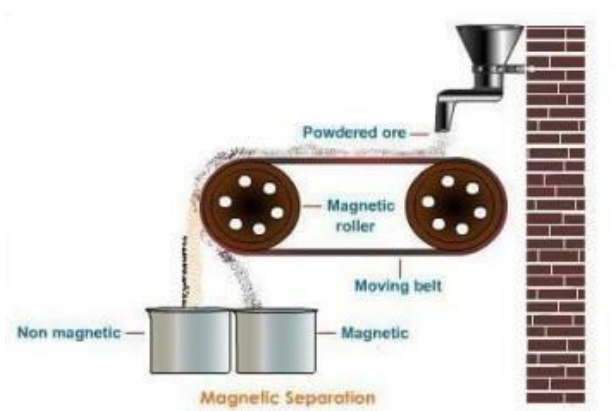
To separate sulphide ores, this technique uses a combination of water and pine oil that is caused to froth in a tank. The ore and gangue particles are distinguished by their different wetting characteristics.



A tank is initially filled with a mixture of water, pine oil, detergent, and powdered ore. To make froth, pressurised air is pushed through the pipe of a revolving agitator. Pine oil wets and coats the sulphide ore particles, which rise up with the froth (froth being lighter). Water-soaked gangue particles settle to the bottom of the tank (water being heavier). The covalent oil molecules are drawn to sulphide because it is more electronegative. The water attracts the gangue because it is less electronegative. The sulphide ore-containing froth is moved to a new container, cleaned, and dried.

Magnetic Separation

This technique uses the difference in magnetic characteristics of the ore and gangue particles to enrich magnetic ores such as pyrolusite (MnO_2) and chromite ($\text{FeO} \cdot \text{Cr}_2\text{O}_3$).

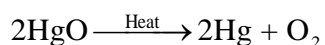
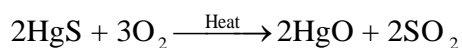


The powdered ore is put onto a conveyor belt that runs between two magnetic rollers. The magnetic ore particles are attracted to the magnetic roller and travel for a bit longer with the conveyor belt than the non magnetic gangue. The gangue particles are the first to fall, producing a heap. The magnetic mineral particles then fall to the ground, producing a distinct heap. As a result, two distinct piles of ore and gangue particles develop.

Extracting Metals Low in the Activity Series

Metals with a low activity series are extremely non-reactive. By simply heating these metals' oxides, they may be reduced to metals. Mercury, for example, is extracted from its ore, cinnabar (HgS), by a heating process.

Reduction using heat



By treating copper sulphide (Cu₂S) ore in the same way, pure copper can be obtained by the similar method as the above method.

Note: Decomposition of mercury and copper oxide formed takes place after the reaction is completed.

Extracting Metals in the Middle of the Activity Series

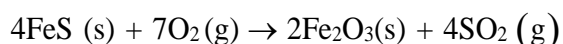
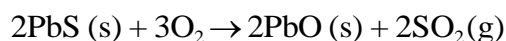
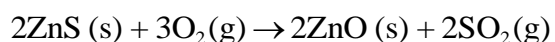
Iron, zinc, lead, copper, and other metals are in the centre of the reactivity series. Sulphides and carbonates are the most common forms of these moderately reactive metals.

Reduction techniques are used to extract these metals from their ores. The oxide ore is reduced during the reduction process.

When compared to sulphides and carbonates, an oxide ore is easier to reduce. If the ore is not an oxide ore, it must first be transformed to an oxide by calcination or roasting.

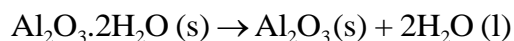
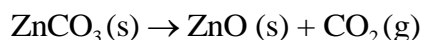
Roasting

Sulphide ores are transformed to oxides by heating them rapidly in the presence of abundant air, causing oxygen to be added, resulting in the formation of the appropriate oxides. Impurities of sulphur escape as gas. Roasting is the term for this procedure.



Calcination

The ore is heated to a high temperature in the absence of air, or when air is not present throughout the reaction. Carbonate ores, as well as ores containing water, are usually calcined to remove carbonate and moisture impurities.



Differences between Roasting and Calcination

Roasting	Calcination
Roasting is the process of heating the ore in the presence of air.	Calcination is the process of heating the ore in the absence of air.

<p>Sulphide ores are roasted in this process. When sulphide ores are roasted, sulphur dioxide as a gas is released.</p> $2\text{ZnS (s)} + 3\text{O}_2\text{(g)} \xrightarrow{\Delta} 2\text{ZnO (s)} + 2\text{SO}_2\text{(g)}$	<p>Carbonate ores are calcinated in this process. Carbon dioxide gas is released when carbonate ores undergo calcination.</p> $\text{CaCO}_3\text{(s)} \xrightarrow{\Delta} \text{CaO (s)} + \text{CO}_2\text{(g)}$
<p>The roasting process is carried out at high temperatures. The temperature sometimes exceeds the melting point of the ore.</p>	<p>The calcinations process is carried out at lower temperatures. The temperature sometimes is below the melting point of the ore.</p>
<p>Roasting is used to eliminate impurities in the form of volatile oxides, oxidise the ore, and remove any moisture that may be present.</p>	<p>Calcination is used to eliminate contaminants, such as volatile oxides, as well as moisture.</p>

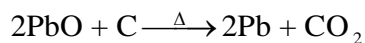
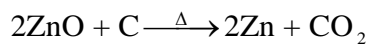
Reduction of Ores

Carbon or hydrogen is used to decrease the oxide produced by calcination or roasting. Coke, or carbon monoxide, is the most common form of carbon. Carbon or carbon monoxide cannot reduce all metallic oxides. Thus, the technique used to extract metal from its ore is determined by the metal's reactivity, or its location in the metal activity series.

Metal	Method of Reduction of Oxide
Zinc (Zn)	Zinc metal can be reduced by carbon reduction method. Only coke can reduce zinc metal.
Iron (Fe), Lead (Pb), Copper (Cu)	The carbon reduction technique is used to reduce these metals. Coke and carbon monoxide can both reduce these metals. Hydrogen can also be used to reduce them.

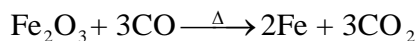
Reduction Using Carbon (Coke)

For oxides of moderately reactive metals, this technique is utilised. Coke is the most commonly used reducing agent since it is affordable.



Reduction Using Carbon Monoxide

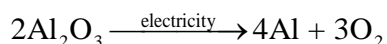
Carbon monoxide is an extremely effective reducing agent. It is used in the blast furnace for hematite reduction.



Extracting Metals towards the Top of the Activity Series

This method is used to make oxides of highly reactive metals that are higher in the reactivity series than aluminium. It can also be used to reduce aluminium oxide.

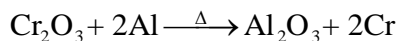
Example:



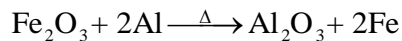
The expense of keeping the electrolyte molten is quite expensive since alumina has a very high melting point. When alumina is combined with cryolite (Na_3AlF_6) and fluorspar (CaF_2), the melting point is significantly reduced, and the cost is also reduced. The conductivity of the electrolyte is increased by these compounds.

Reduction Using Aluminium (Aluminothermy)

Aluminum is a rather costly reducing agent. It is employed in the reduction of extremely reactive metal oxides. However, in the reactivity sequence, these metals must be below aluminium. Example:



Welding damaged iron rails is also done with aluminothermy. The heat created in the reaction causes the metallic iron synthesised to melt since the process is very exothermic. As the molten iron cools, it drops down onto the two pieces to be welded, joining them together.



Refining of Metals

The majority of metals acquired during the reduction process are not very pure. These must be processed or purified further. The final phase in metallurgy is metal purification. The difference between the characteristics of metals and their impurities is used in refining. For refining, we utilise the following procedure.

Electro refining

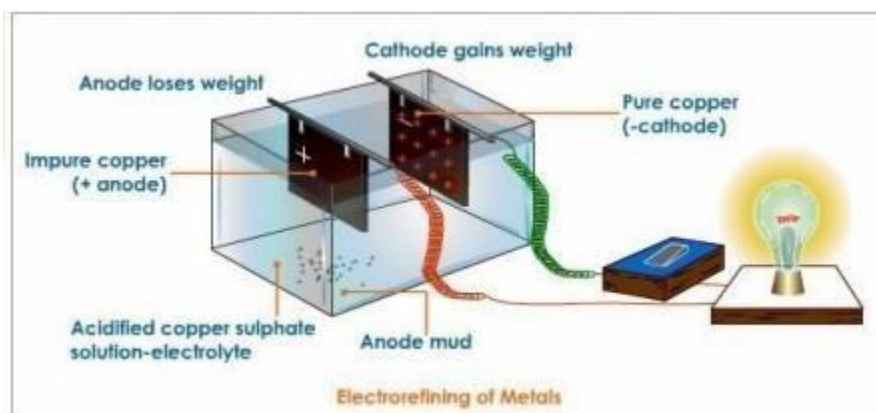
Electrolysis can be used to recover metals that cannot be separated via a chemical reduction technique, as well as to purify metals acquired through other means. The anode in the electro refining process is a block of impure metal, while the cathode is a thin sheet of pure metal. An aqueous solution of the metal salt is included in the electrolytic cell. When an electric current of a sufficient voltage is passed via the anode, impure metal is dissolved and pure metal is deposited at the cathode. The following is how metal ions from the anode enter the electrolyte:



These ions get deposited on the cathode in the following manner



Near the anode, the contaminants are left as anode mud. Due to the accumulation of pure metal, the anode eventually disintegrates, while the cathode accumulates weight.



This technique is used to refine volatile metals with lower boiling points than their impurities, such as copper, silver, tin, and nickel. For example: Mercury and Zinc.

- An electrolyte is a substance (salt, acid, or base) that transmits an electric current in solution or in a molten form while also being decomposed by it. The current is carried by ionised electrolytes, which are electrically charged ions.
- Charged ions migrate towards oppositely charged electrodes in order to lose their electric charge and form atoms, which are then either released or deposited at the electrodes.

Corrosion of Metals

We've learned that in the presence of a wet environment, chemically active metals corrode. Corrosion is an oxidation process involving ambient oxygen and water on a metal's surface. Iron corrodes faster than most other transition metals, resulting in the formation of iron oxide. In the presence of carbon dioxide, as well as salt solution, the corrosion or rusting of iron is increased.

Rusting is: $2\text{Fe (s)} + \text{O}_2\text{(g)} + \text{H}_2\text{O (l)} \rightarrow \text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O (s)}$

Rust is also known as hydrated iron oxide.

Rusting is an oxidation reaction because it includes iron atoms losing electrons or iron atoms obtaining oxygen. The amount of water 'x' is variable, ranging from dry to wet, and the equation is not supposed to be balanced.

Corrosion is a serious issue in sectors that utilise iron (or steel) as a structural material, such as construction, infrastructure, bridges, rail transport power transmission, shipbuilding, cars, and heavy industries.

Aluminum, another valuable structural metal, undergoes an oxidation process as well, but not as fast as its reactivity indicates. Once a small coating of Al_2O_3 has developed on the surface, it acts as a barrier to oxygen and water, preventing the aluminium from corroding further. As a result, aluminium is known as a self-protective metal. Aluminum alloys may be used to make it harder, stronger, and stiffer by combining it with tiny quantities of other metals (such as magnesium).

Alkali metals, such as sodium, which are utilised in chemistry laboratories and some chemical industries, corrode quickly in the air and must be kept under oil.

Copper and lead are both utilised in roofing because they are both non-reactive. The surface chemicals do not peel away as readily as rust does off iron. Copper corrodes to create a basic green carbonate (combination of the hydroxide and the carbonate), while lead corrodes to form a white lead oxide or carbonate. Both metals have been used for pipes in the past, but because lead is deemed too hazardous, copper is now the most common choice. Gold, platinum, and mercury are non-reactive metals that do not corrode.

Prevention of Corrosion

Paint, which creates a barrier between the metal and air/water, is the best way to preserve iron and steel (alloys of iron). A water resistant oil or grease coating can be used to protect moving components on machinery. Another option is to use enamel and lacquers to cover the surface.

Sacrificial Protection

By linking iron to a more reactive metal, 'rusting' can be avoided (e.g., zinc or magnesium). Because the more reactive shielding metal is preferentially oxidised away, leaving the protected metal intact, this is referred to as sacrificial protection or sacrificial corrosion.

Alloying

Iron and steel, as well as other metals, may be made non-rusting alloys by 'alloying,' or combining with other metals (e.g., chromium). Stainless steel is an iron and carbon alloy that does not rust. Brass, a copper-based alloy, is another less costly and non-reactive metal alloy.

Galvanizing

The process in which iron or steel is coated with a thin layer of zinc is called galvanizing. Electrolytic deposition is used to create this layer. The negative cathode zinc is deposited on the iron/steel item after it is dipped in molten zinc. Zinc corrodes or oxidises preferentially, forming a zinc oxide coating that does not flake off like iron oxide rust does. In addition, if the surface is damaged, the exposed zinc corrodes first, protecting the iron.

Electroplating

Electroplating is also used to prevent corrosion by coating the surface with metals such as tin, chromium, and nickel. Steel cans are protected by a thin layer of unreacted tin that lasts as long as the tin layer is complete.