Chapter 3 - Metals and Non-metals

Introduction

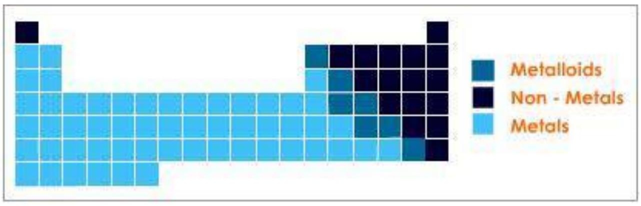
There are 118 elements present in the periodic table, 92 of which are naturally occurring. Metals and non-metals are characterized by distinctly different physical and chemical properties. At present about 80 metals are known to us.

At room temperature, over half of the non-metals are gases, except bromine, which is a liquid.

The most abundant non-metal in the earth's crust is oxygen, which constitutes about 50% of the earth's crust and along with nitrogen it forms the main constituents of air.

The next abundant nonmetal is silicon which constitutes about 26% of the earth's crust. Oxygen and silicon are the two major constituents of earth. Hydrogen and oxygen are the two major constituents of the oceans.

Position of Metals and Non-metals in the Periodic Table



Metals occupy the groups on the left of the periodic table. Group IA consists of highly reactive metals called the alkali metals, while group II A elements are called alkaline earth metals. Elements between group IIA and IIIA are all called transition metals.

The non-metals are elements (with the exception of hydrogen) that are found to the right on the Periodic Table i.e., groups IVA, VA, VIA &VIIA. The non-metallic character of these elements increases from top to the bottom of the group. For example, in group VA the first and second members are non-metals, the third and fourth are metalloids and the last member is a metal. The metalloids are a group of elements which have properties similar to both the metals and non-metals. These metalloids are: Boron, silicon, germanium, arsenic, antimony, tellurium and astatine. The non-metals are elements found to the right of these metalloids, including the element, hydrogen.

Group V A	Non Metals	Nitrogen, Phosphorous,
	Metalloids	Arsenic, Antimony
	Metal	Bismuth

Physical Properties of Metals

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

Lustre - Metals have the quality of reflecting light from its surface and can be polished e.g., gold, silver and copper.

Malleability - Metals have the ability to withstand hammering and can be made into thin sheets known as foils. Except Zinc which is brittle.

Ductility - Metals can be drawn into wires. Except Zinc which is brittle.

Hardness - All metals are hard except sodium and potassium, which are soft and can be cut with a knife.

Conduction - Metals are good conductors because they have free electrons. Silver and copper are the two best conductors of heat and electricity. Lead is the poorest conductor of heat. Bismuth, mercury and iron are also poor conductors

Density - Metals have high density and are very heavy. Iridium and osmium have the highest densities whereas lithium has the lowest density.

Melting and Boiling Point - Metals have high melting and boiling point. Tungsten has the highest melting point where as silver has low boiling point. Sodium and potassium have low melting points.

Alloy Formation - Metals form homogeneous mixture with each other called an alloy. Example- Brass is an alloy of copper and zinc.

Sonorous - Metals are sonorous i.e. they produce sound when hit with some solid object.

Physical Properties of Non-metals

Physical State - Most of the non-metals exist in two of the three states of matter at room temperature: gases (oxygen) and solids (iodine, carbon, sulphur). These have no metallic lustre, (except iodine) and do not reflect light. (Except carbon in the form of diamond).

Nature - Non-metals are very brittle, and cannot be rolled into wires or pounded into sheets. Except- diamond is the hardest substance known.

Conduction - They are poor conductors of heat and electricity. (Except graphite conducts heat, both graphite & gas carbon conduct electricity.)

Electronegative Character - Non-metals have a tendency to gain or share electrons with other atoms. They are electronegative in character.

Reactivity - They generally form acidic or neutral oxides with oxygen.

Melting and Boiling Points – Non-metals have low melting and boiling points.

Comparative Properties of Metals and Non-Metals

Property	**Metals**	**Non-metals**
State of matter	These are usually solid, except mercury,	These exist in all the three states.

Property	**Metals**	**Non-metals**
	which is a liquid at room temperature. Gallium and Caesium melt below 30^{o}C. So if room temperature is around 30^{o}C, they may also be in liquid state	Bromine is the only liquid. Solids – iodine, carbon, sulphur.
Density	They usually have high density, except for sodium, potassium, calcium etc.	Their densities are usually low.
Melting point	They usually have a high melting point except mercury, cesium, gallium, tin, lead.	Their melting points are low.
Boiling point	Their boiling points are usually high.	Their boiling points are low.
Hardness	They are usually hard, except mercury, sodium, calcium, potassium, lead etc.	They are usually not hard. But the exception is the non-metal diamond, the hardest substance.
Malleability	They can be beaten into thin sheets.	They are generally brittle.
Ductility	They can be drawn into thin wires, except sodium, potassium, calcium etc.	They cannot be drawn into thin wires.
Conduction of heat	They are good conductors of heat.	They are poor conductors of heat. (exception- carbon in the form of graphite)
Conduction of electricity	They are good conductors of electricity.	They are non-conductors, except for carbon in the form of graphite and the gas carbon.
Lustre	Newly cut metals have high lustre. Some get tarnished immediately.	Usually not lustrous, except iodine and diamond - the most lustrous of all the substances.
Alloy formation	They form alloys.	Generally, they do not form alloys. However, carbon, phosphorus, sulphur etc. can be present in some alloys.
Tenacity	These usually have high tensile strength except sodium, potassium, calcium, lead etc.	These have low tensile strength.
Brittleness	They are hard but not brittle, except zinc at room temperature.	They are generally brittle.
Electronic	They usually have 1, 2 or 3 electrons in their valence shell. The greater the	They usually have 4, 5, 6 or 7 electrons in the valence shell. If it

Property	**Metals**	**Non-metals**
configuration	number of shells and lesser the number of valence electrons, the greater is the reactivity of the metal.	has 8 electrons, it is called a noble gas. Lesser the number of shells and greater the number of valence electrons, greater is the reactivity of the non-metal.
Ionization	They always ionize by losing electrons: $Na^0 - e^- \rightarrow Na^+$	They always ionize by gaining electrons: $Cl^0 + e^- \rightarrow Cl^-$
Charge of ions	Positively charged.	Negatively charged.
Type of valency	Metals always exhibit electrovalency.	Non-metal exhibit both electrovalency or covalency.
Deposition during electrolysis	They are always deposited at the cathode.	They are always deposited at the anode.
Redox reaction	These lose electrons and hence get oxidized.	These gain electrons and hence get reduced.
Redox agents	They are reducing agents.	They are oxidizing agents.
Nature of oxides	They generally form basic oxides, some of which are also amphoteric, such as aluminum oxide, zinc oxide, lead oxide etc.	They generally form acidic oxides. Neutral oxides are nitrous oxide, nitric oxide, carbon monoxide water etc.
Hydrides	They do form hydrides except some transition elements.	They do form hydrides, e.g. NH_3, PH_3, HCl, HBr, HI, H_2S, H_2O etc.
Atomicity	These are always monatomic.	These can be mono, di, tri, or polyatomic.
Solubility	They do not dissolve in solvents except by chemical action.	They dissolve in solvents and can be re-obtained by evaporation. Example: Sulphur in carbon disulphide.
Action with chlorine	They produce chlorides, which are electrovalent.	They produce chlorides, which are covalent.
Action with dilute acids	On reaction with dilute acids they give respective salt and hydrogen.	They do not react with dilute acids.

Chemical Properties of Metals

Metals are Electropositive Elements

Metals are very reactive. Metals tend to lose electrons easily and form positively charged ions; therefore, metals are called electropositive elements. Sodium metal forms sodium ions Na^+. The electropositive nature allows metals to form compounds with other elements easily.

Reaction of Metals with Oxygen

Metals like sodium (Na) and potassium (K) are some of the most reactive metals. Potassium, sodium, lithium, calcium and magnesium react with oxygen and burn in air.

Metals from aluminum to copper in the activity series of metals react slowly when heated in air to form the metal oxides. Aluminum is the fastest and copper is the slowest of them.

- Sodium metal reacts with the oxygen of the air at room temperature to form sodium oxide. Hence, sodium is stored under kerosene to prevent its reaction with oxygen, moisture and carbon dioxide.

$$\underset{\text{Sodium}}{4Na} + \underset{\text{Oxygen}}{O_2} \longrightarrow \underset{\text{Sodium oxide}}{2Na_2O}$$

- Sodium oxide is a basic oxide which reacts with water to form sodium hydroxide.

$$\underset{\text{Sodium oxide}}{Na_2O} + H_2O \longrightarrow \underset{\text{Sodium hydroxide}}{2NaOH}$$

- Mg does not react with oxygen at room temperature. On heating, Mg burns in air with intense light and heat to form MgO.

$$2Mg + O_2 \xrightarrow{\Delta} 2MgO$$

- Zinc metal burns in air only on strong heating to form zinc oxide.

$$\text{Zinc} + \text{Oxygen} \xrightarrow{\Delta} \text{Zinc oxide}$$

$$2Zn + O_2 \xrightarrow{\Delta} 2ZnO$$

- In moist air, iron is oxidized to give rust.

$$\text{Iron} + \text{Moist air} \longrightarrow \text{Iron (II, III) oxide}$$

$$3Fe + 2O_2 + xH_2O \longrightarrow Fe_3O_4.xH_2O$$

- On heating in air it burns with a brilliant flame forming triferric tetroxide.

$$3Fe + 2O_2 \xrightarrow{\Delta} Fe_3O_4$$

- Copper is the least reactive metal and does not burn in air even on heating. However, on prolonged strong heating copper reacts with oxygen and forms copper (II) oxide (CuO) outside and copper (I) oxide (Cu_2O) inside.

$$\text{Copper} + \text{Oxygen} \xrightarrow{\Delta} \text{Copper (II) oxide}$$

$$2Cu + O_2 \xrightarrow{\Delta} 2CuO$$

- Gold and platinum do not react with oxygen in air.

Reaction of Metals with Water

Potassium, sodium, lithium and calcium react with cold water.

- Sodium reacts vigorously with cold water forming sodium hydroxide and hydrogen.

  $2Na + H_2O \longrightarrow 2NaOH + H_2 \uparrow$

- Metals from magnesium to iron in the activity series of metals, react with steam (but not cold H_2O) to form the metal oxide and hydrogen gas.

  $\text{Magnesium} + \text{Steam} \longrightarrow \text{Magnesium oxide} + \text{Hydrogen}$

  $Mg + H_2O \longrightarrow MgO + H_2 \uparrow$

  $\text{Aluminium} + \text{Steam} \longrightarrow \text{Aluminium oxide} + \text{Hydrogen}$

  $2Al + 3H_2O \longrightarrow Al_2O_3 + 3H_2 \uparrow$

  $\text{Zinc} + \text{Steam} \longrightarrow \text{Zinc oxide} + \text{Hydrogen}$

- Red hot iron reacts with steam to form Iron (II, III) oxide.

  $\text{Iron} + \text{Steam} \Leftrightarrow \text{Iron (III) oxide} + \text{Hydrogen}$

  $3Fe + 4H_2O \Leftrightarrow Fe_3O_4 + 4H_2 \uparrow$

  $2Fe + 3H_2O_{(g)} \Leftrightarrow Fe_2O_3 + 3H_2 \uparrow$

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

- Potassium, sodium, lithium and calcium react violently with dilute H_2SO_4 and dilute HCl, forming the metal salt (either sulphate or chloride) and hydrogen gas. The reaction is similar to the reaction with water.

  $\text{Sodium} + \text{Hydrochloric acid} \longrightarrow \text{Sodium chloride} + \text{Hydrogen}$

  $2Na + 2HCl_{(dil)} \longrightarrow 2NaCl + H_2 \uparrow$

- Magnesium, aluminum, zinc, iron, tin and lead react safely with dilute acid. Magnesium is the fastest and lead is the slowest of the six.

  $\text{Magnesium} + \text{Sulphuric acid} \longrightarrow \text{Maganesium sulphate} + \text{Hydrogen}$

  $Mg + H_2SO_4 \longrightarrow MgSO_4 + H_2 \uparrow$

  $\text{Magnesium} + \text{Hydrochloric acid} \longrightarrow \text{Magnesium chloride} + \text{Hydrogen}$

  $Mg + 2HCl \longrightarrow MgCl_2 + H_2 \uparrow$

  $\text{Zinc} + \text{Sulphuric acid} \longrightarrow \text{Zinc sulphate} + \text{Hydrogen}$

  $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2 \uparrow$

  $\text{Iron} + \text{Hydrochloric acid} \longrightarrow \text{Iron (II) chloride} + \text{Hydrogen}$

  $Fe + 2HCl \longrightarrow FeCl_2 + H_2 \uparrow$

Zinc with dilute sulphuric acid is often used for the laboratory preparation of hydrogen. The reaction is slow at room temperature, but its rate can be increased by the addition of a little copper (II) sulphate. Zinc displaces copper metal, which acts as a catalyst.

Metals below hydrogen (copper, silver, gold and platinum), will not react with dilute acid to liberate hydrogen. In general,

- Hydrochloric acid makes a metal chloride.
- Sulphuric acid makes a metal sulphate.
- Reactions with nitric acid are more complex, the nitrate is formed but the gas is rarely hydrogen, and more often, an oxide of nitrogen.

Reaction of Metals with Salt Solutions

Reactive metals can displace any metal less reactive than itself, from the oxide, chloride or sulphate of the less reactive metal in solution or their molten state. If metal A displaces metal B from its solution, it is more reactive than B.

$$\begin{array}{lllllll} \text{Metal A} & + & \text{Salt solution of B} & \longrightarrow & \text{Salt solution of A} & + & \text{Metal B} \\ \text{Iron} & + & \text{Copper (II) sulphate} & \longrightarrow & \text{Iron (II) sulphate} & + & \text{Copper} \\ Fe & + & CuSO_4 & \longrightarrow & FeSO_4 & + & Cu \end{array}$$

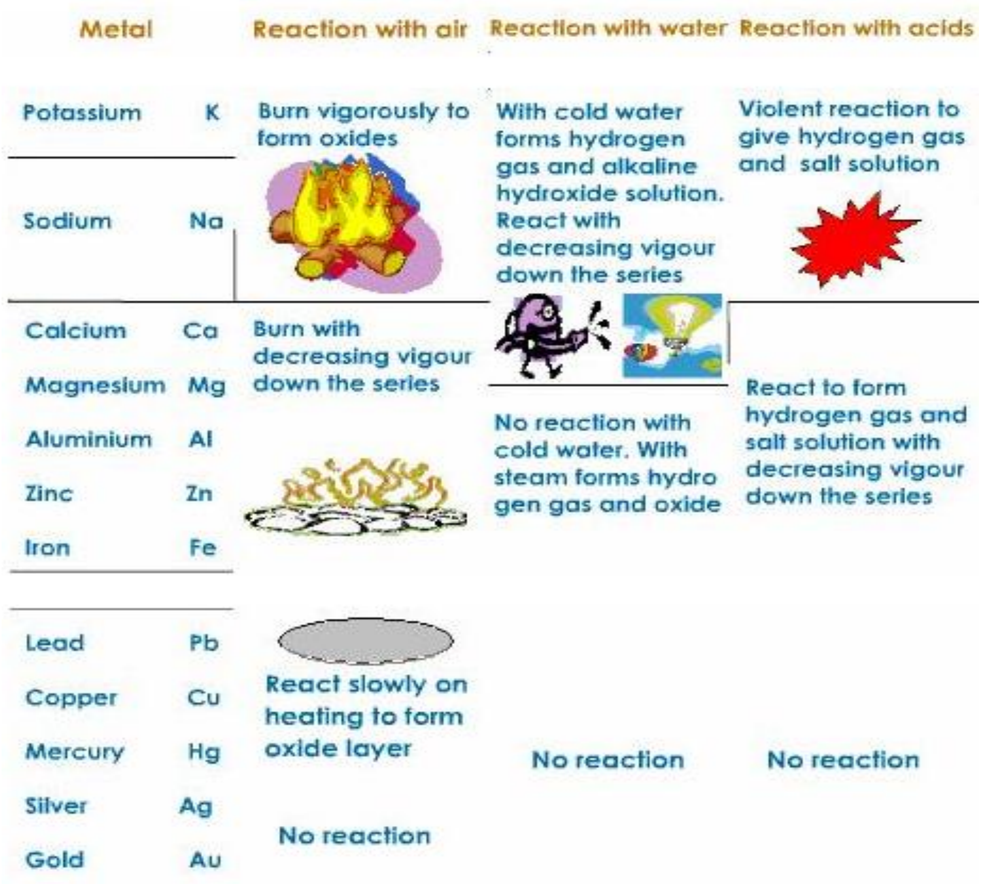
Copper (II) sulphate solution is blue; iron sulphate solution is almost colourless when dilute. During the displacement, the blue solution loses its color and the iron metal is seen to turn pink-brown as the displaced copper becomes deposited on it.

On heating the mixture of magnesium powder and black copper (II) oxide, white magnesium oxide is formed with brown bits of copper:

$$Mg + CuO \xrightarrow{\Delta} MgO + Cu$$

Adding magnesium to blue copper (II) sulphate solution, the blue color fades as colourless magnesium sulphate is formed and brown bits of copper metal form a precipitate:

$$Mg + CuSO_4 \Rightarrow MgSO_4 + Cu$$



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		React to form hydrogen gas and salt solution with decreasing vigour down the series
Magnesium	Mg		No reaction with cold water. With steam forms hydro gen gas and oxide	
Aluminium	Al			
Zinc	Zn			
Iron	Fe			
Lead	Pb	React slowly on heating to form oxide layer	No reaction	No reaction
Copper	Cu			
Mercury	Hg			
Silver	Ag	No reaction		
Gold	Au			

Electronic Nature of Metals and Non-metals

The atoms of all elements, except noble gases, have an incomplete outermost shell. . Noble gases have their outermost shell complete and hence they are not reactive or "inert".

Most elements are reactive and try to achieve the stability of the noble or inert gases by electron transfer or by electron sharing. Elements that can donate electrons are called metals. They form positive ions by losing electrons.

The elements that accept electrons are called non-metals. They form negative ions by gaining electrons. Metals have 1 to 3 electrons in the outermost shell of their atom and non-metals have 4 to 8 electrons in the outermost shell.
There are two exceptions to this rule: Hydrogen and helium. Hydrogen is a non-metal having 1 electron in the valence shell and helium too is an inert gas having 2 electrons in the valence shell.

Type of Elements	**Element**	**Atomic Number**	**Number of Electrons in Shells**			
			K	L	M	N
Noble Gases	Helium (He)	2	2			
	Neon (Ne)	10	2	8		

Type of Elements	**Element**	**Atomic Number**	**Number of Electrons in Shells**			
	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		
	Phosphorus (P)	15	2	8	5	
	Sulphur (S)	16	2	8	6	
	Chlorine (Cl)	17	2	8	7	

Important Point

Metals that donate electrons gain positive charge equal to the number of electrons donated. For example, atomic number of aluminum is 13, so the electronic configuration of Al is 2, 8, and 3. Aluminum has 3 electrons in the valence shell; it loses 3 electrons to form Al^{3+}.

$$Al^{o} - 3e^{-} \longrightarrow Al^{3+}$$

Other examples,

$$Na^{0} - 1e^{-} \longrightarrow Na^{+}$$

$$Mg^{0} - 2e^{-} \longrightarrow Mg^{2+}$$

Non-metals gain electrons and hence gain negative charge equal to the number of electrons accepted.

For example,

$$Cl^{0} + 1e^{-} \longrightarrow Cl^{-}$$

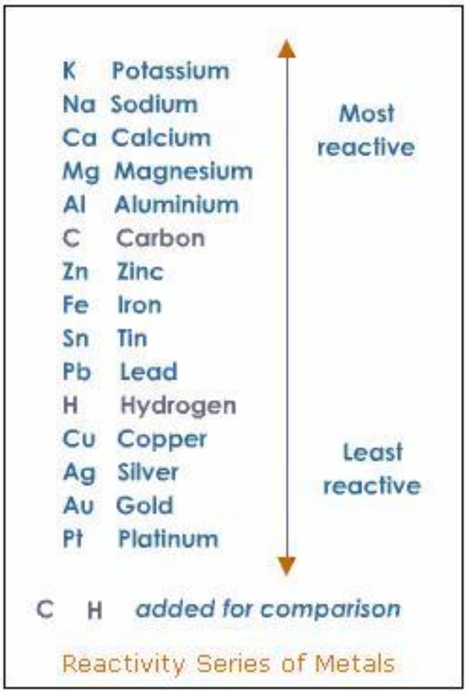
$$S^{0} + 2e^{-} \longrightarrow S^{2-}$$

$$P^{0} + 3e^{-} \longrightarrow P^{3-}$$

The Reactivity Series of Metals

Although most metals are usually electropositive in nature and lose electrons in a chemical reaction they do not react with the same vigour or speed. Metals display different reactions towards different substances. The greater the ease with which an element loses its electrons and acquires a positive charge, the greater is its reactivity. Further, the greater the number of shells and lesser the number of valence electrons, the greater is the reactivity of the metal. The activity series of metals, arranges all metals in order of their decreasing chemical activity. As we go down the activity series from potassium to gold the ease with which a metal loses electrons, and forms positive ions in solutions, decreases.

The most active metal, potassium, is at the top of the list and the least reactive metal, gold, is at the bottom of the list. Although hydrogen is a non-metal it is included in the activity series due to the fact that it behaves like a metal in most chemical reactions i.e., the hydrogen ion has a positive charge [H^{+}] like other metals.



Reactivity Series of Metals

Element	Symbol	Group Number
Potassium	K	IA
Sodium	Na	IA
Lithium	Li	IA
Calcium	Ca	IIA
Magnesium	Mg	IIA
aluminium	Al	IIIA
Carbon	C	IVA
Zinc	Zn	IIB
Iron	Fe	VIII
Tin	Sn	IVA
Lead	Pb	IVA
Hydrogen	H	IA
Copper	Cu	IB
Silver	Ag	IB

Element	Symbol	Group Number
Gold	Au	IB
Platinum	Pt	VIII

- The higher the metal in the series, the more reactive it is i.e., its reaction is fast and more exothermic.
- This also implies that the reverse reaction becomes more difficult i.e., the more reactive a metal, the more difficult it is to extract from its ore. The metal is also more susceptible to corrosion with oxygen and water.
- The reactivity series can be established by observation of the reaction of metals with water, oxygen or acids.
- Within the general reactivity or activity series there are some periodic table trends:

Metals	Reactivity and reactions
Potassium K	Very reactive, very rapid with cold water forming the alkali potassium hydroxide and hydrogen gas (which is ignited). $2K+2H_2O_{(l)} \Rightarrow 2KOH_{(aq)} + H_{2(g)}$
Sodium Na	Fast reaction with cold water forming the alkali sodium hydroxide and hydrogen gas. $2Na_{(s)} + 2H_2O_{(l)} \Rightarrow 2NaOH_{(aq)} + H_{2(g)}$ The reaction of sodium with water-the sodium melts to a silvery ball and fizzes as it spins over the water. The rapid exothermic reaction produces a colourless gas that gives a squeaky pop! with a lit splint-hydrogen. Universal indicator will turn from green to purple/violet-the strong alkali sodium hydroxide is formed. The sodium floats because it is less dense than water.
Calcium Ca	Quite reactive with cold water forming the moderately soluble alkali calcium hydroxide and hydrogen gas. $Ca_{(s)} + 2H_2O_{(l)} \Rightarrow Ca(OH)_{2(aq/s)} + H_{2(g)}$ Very reactive with dilute hydrochloric acid forming the colourless soluble salt calcium chloride and hydrogen gas. $Ca_{(s)} + 2HCl_{(g)} \Rightarrow CaCl_{2(aq)} + H_{2(g)}$ Not very reactive with dilute sulphuric acid because the colourless calcium sulphate formed is not very soluble and coats the metal inhibiting the reaction. $Ca_{(s)} + H_2SO_{4(aq)} \longrightarrow CaSO_{4(s)} + H_{2(g)}$

Metals	Reactivity and reactions
Magnesium Mg	Slow reaction with water forming the slightly soluble base magnesium oxide and hydrogen gas. With steam, the reaction is faster with heated magnesium and a white powder magnesium oxide is formed along with hydrogen. Magnesium will burn with a bright white flame in steam, if previously ignited in air. $Mg_{(s)} + H_2O_{(g)} \longrightarrow MgO_{(s)} + H_{2(g)}$ In fact it will even burn in carbon dioxide forming black specks of carbon! $2Mg_{(s)} + CO_{2(g)} \longrightarrow 2MgO_{(s)} + C_{(s)}$ Very reactive with dilute hydrochloric acid forming the colourless soluble salt, magnesium chloride and hydrogen gas. $Mg_{(s)} + 2HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_{2(g)}$ Very reactive with dilute sulphuric acid forming colourless soluble magnesium sulphate and hydrogen. $Mg_{(s)} + H_2SO_{4(aq)} \longrightarrow MgSO_{4(aq)} + H_{2(g)}$
Aluminum Al	Aluminum has no reaction with water or steam due to a protective aluminum oxide layer of Al_2O_3. Slow reaction with dilute hydrochloric acid to form a colourless soluble salt aluminum chloride and hydrogen gas. $2Al_{(s)} + 6HCl_{(aq)} \Rightarrow 2AlCl_{3(aq)} + 3H_{2(g)}$ The reaction with dilute sulphuric acid is extremely slow to form colourless aluminum sulphate and hydrogen. $2Al_{(s} + 3H_2SO_{4(aq)} \Rightarrow Al_2(SO_4)_{3(aq)} + 3H_{2(g)}$
(Carbon C,a non-metal)	Elements higher than carbon i.e aluminum and the more reactive metals must be extracted by electrolysis (or displacing it with an even more reactive metal). Metals below it, i.e., zinc or a less reactive can be extracted by reducing the hot metal oxide with carbon.
Zinc Zn	No reaction with cold water. When the metal is heated in steam zinc, oxide and hydrogen are formed. $Zn_{(s)} + H_2O_{(g)} \Rightarrow ZnO_{(s)} + H_{2(g)}$ Quite reactive with dilute hydrochloric acid forming the colourless soluble salt zinc chloride and hydrogen gas. $Zn_{(s)} + 2HCl_{(aq)} \Rightarrow ZnCl_{2(aq)} + H_{2(g)}$ Quite reactive with dilute sulphuric acid forming the colourless soluble salt zinc sulphate and hydrogen gas.

Metals	Reactivity and reactions
	$Zn_{(s)} + H_2SO_{4(g)} \Rightarrow ZnSO_{4(s)} + H_{2(g)}$ (this reaction is catalyzed by adding a trace of copper sulphate solution) Zinc can be extracted by reducing the hot metal oxide on heating with carbon $2ZnO_{(s)} + C_{(s)} \Rightarrow 2Zn_{(s)} + CO_{2(g)}$ A zinc coating (galvanizing) is used to protect iron from rusting.
Iron Fe	No reaction with cold water (rusting is a joint reaction with oxygen). When the metal is heated in steam an iron oxide (unusual formula) and hydrogen are formed. This oxide is 'technically' Iron (III, II) oxide! $3Fe_{(s)} + 4H_2O_{(g)} \Rightarrow Fe_3O_{4(s)} + 4H_{2(g)}$ Slow reaction with dilute hydrochloric acid forming the soluble pale green salt Iron (II) chloride and hydrogen gas. $Fe_{(s)} + 2HCl_{(aq)} \Rightarrow FeCl_{2(aq)} + H_{2(g)}$ Slow reaction with dilute sulphuric acid forming the soluble pale green salt Iron (II) sulphate and hydrogen gas. $Fe_{(s)} + 2H_2SO_{4(g)} \Rightarrow FeSO_{4(s)} + H_{2(g)}$ Iron can be extracted by reducing the hot metal oxide on heating with carbon monoxide formed from carbon in the blast furnace e.g., $Fe_2O_{3(s)} + 3CO_{(g)} \Rightarrow 2Fe_{(s)} + 3CO_{2(g)}$ $Fe_3O_{4(s)} + 4CO_{(g)} \Rightarrow 3Fe_{(s)} + 4CO_{2(g)}$
(Hydrogen H non-metal)	None of the metals below hydrogen can react with acids to form hydrogen gas. They are least easily corroded metals and partly accounts for their value and uses in jewellery, electrical contacts etc.
Copper Cu	No reaction with cold water or when heated in steam. No reaction with dilute hydrochloric acid or dilute sulphuric acid. Copper can be extracted by reducing the hot black metal oxide on heating with carbon. $2CuO_{(s)} + C_{(s)} \Rightarrow 2Cu_{(s)} + CO_{2(g)}$
Silver Ag	No reaction with cold water or when heated in steam. No reaction with dilute hydrochloric acid or dilute sulphuric acid. Silver can be extracted by reduction but can be found 'native' as the element.
Gold Au	No reaction with cold water or when heated in steam. No reaction with dilute hydrochloric acid or dilute sulphuric acid. Gold can be readily extracted from its ores easily by reduction but it is usually found 'native'. Pure gold is 24 carat.

Metals	Reactivity and reactions
Platinum Pt	No reaction with cold water or when heated in steam. No reaction with dilute hydrochloric acid or dilute sulphuric acid. Like gold, it is a very rare metal. It is used in expensive jewellery, laboratory ware (inert crucible container) as a industrial catalyst, and catalytic converters in car exhausts.

Bonding

The tendency of an atom to take part in chemical combination is determined by the number of valence electrons (electrons in the outermost shell of an atom). The atoms acquire the stable noble gas configuration of having eight electrons in the outermost shell (called octet rule) during chemical combination.

The combination of atoms occurs in two ways: either by electrovalent bonding or covalent bonding. In all chemical reactions, it is the electrons from the outermost shell of an atom that are involved in interacting with other atoms, either by their transfer or by sharing.

Electrovalent Bonding

When an atom donates one, two or three electrons from its valence shell to another atom, which has the ability to accept these electrons, it is known as electrovalency. As a result of electrovalency, both these atoms achieve the structure of an inert gas. When the chemical bond occurs by the transfer of electrons from the atom of an element to the atom or atoms of another it is called Ionic or Electrovalent bond.

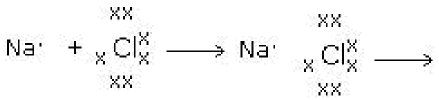
Thus, the electrovalency of sodium is 1+, and that of chlorine is 1- in NaCl. Similarly, calcium, magnesium in their chloride exhibits an electrovalency of 2+. There are many elements, which show different electrovalencies in different compounds. This phenomenon is called 'variable electrovalency' e.g., iron exists as Fe^{2+} and Fe^{3+} in ferrous sulphate and ferric sulphate respectively.

Formation of Sodium Chloride

During the formation of an ionic bond between the metal sodium and the non-metal chlorine, sodium loses one electron to complete its octet as it has only one electron in its valence shell. It acquires a noble gas configuration of neon (2, 8). While the chlorine atom has seven electrons in its valence shell and gains one electron to complete its octet and also acquires stable electronic configuration of argon.

$$\underset{\text{Sodium atom}}{\underset{(2,8,1)}{Na}} \xrightarrow{\text{lossot electron}} \underset{\text{sodium ion (neon configuration)}}{\underset{(2,8)}{Na^+}} + \text{electron}$$

$$\underset{\text{Chlorine atom}}{\underset{(2,8,7)}{Cl}} + \text{electron} \xrightarrow{\text{gain of electron}} \underset{\text{chloride ion (argon configuration)}}{\underset{(2,8,8)}{Cl^-}}$$



$$\longrightarrow Na^+ Cl^- + \text{energy}$$

Formation of Magnesium Chloride

Magnesium, whose atomic number is 12, has 2, 8, 2 configurations. Its valence shell has two electrons. The electronic configuration of chlorine (At. no. 17) is 2, 8, 7. It has seven valence electrons. Since, magnesium has two electrons in excess of the neon configuration (2, 8), and chlorine is one electron short of the argon configuration (2,8,8), hence one atom of magnesium will look for two atoms of chlorine to transfer its two electrons to (one to each) as shown below:

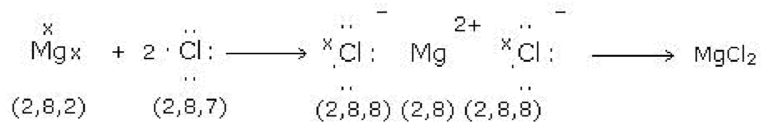
$$\underset{(2,8,2)}{Mg} \longrightarrow \underset{(2,8)}{Mg^{2+}} + 2e^-$$

$$\underset{(2,8,7)}{2Cl} + 2e^- \longrightarrow \underset{(2,8,8)}{2Cl^-}$$

The Mg^{2+} and the two Cl^- so formed, then form ionic bonds between them.

$$Mg^{2+} + 2Cl^- \longrightarrow [Cl^- - Mg^{2+} - Cl^-] \longrightarrow MgCl_2$$

In terms of Lewis dot structure,



Formation of Magnesium Oxide

Mg (at no 12) has configuration 2, 8, 2.

The atom of Mg loses 2 electrons to become stable like Neon (2, 8)

$Mg – 2e^- \rightarrow Mg^{2+}$

Oxygen (at no 8) has configuration 2,6.The atom of Oxygen gains 2 electrons to become stable like Ne (2, 8)

$O + 2e^- \rightarrow O^{2-}$

$Mg^{2+} + O^{2-} \rightarrow MgO$

Formation of Calcium oxide

Ca (at no 20) has configuration 2,8,8,2.

The atom of Ca loses 2 electrons to become stable like Argon (2,8,8)

$Ca – 2e^- \rightarrow Ca^{2+}$

Oxygen (at no 8) has configuration 2, 6.

The atom of Oxygen gains 2 electrons to become stable like Ne (2, 8)

$O + 2e^- \rightarrow O^{2-}$

$Ca^{2+} + O^{2-} \rightarrow CaO$

Properties of Electrovalent Compounds

Property	Electrovalent or ionic compounds
Structure of charged ions	They consist of oppositely charged molecules.
Physical state and hardness	The inter-atomic attraction is high, hence they are brittle, hard, crystalline solids.
Melting and boiling points	Due to strong attraction between the particles, high temperatures are required to melt or boil them.
Solubility	They are usually soluble in water, but insoluble in organic solvents.
Passage of electricity	Ionic compounds do not conduct electricity in the solid state because movement of ions in the solid is not possible due to their rigid structure. In the molten form or in aqueous solution form, since the electrostatic forces of attraction between the oppositely charged ions are overcome they allow the flow of electricity, and get decomposed by it.
Rate of reaction	Their reaction usually occurs with high speeds.
Dissociation in solution	Since electrovalent compounds are made up of charged ions, they dissociate to give negative and positive ions in solution. $NaCl \longrightarrow Na^+ + Cl^-$
Electrolysis	These compounds can undergo electrolysis. The cations get discharged at cathode and anions at anode.

Covalent Bond

A covalent bond is defined 'as the force of attraction arising due to mutual sharing of electrons between the two nonmetallic atoms'. The combining atoms may share one, two or three pairs of electrons. The covalent bond is formed between two similar or dissimilar atoms of nonmetals by a mutual sharing of electrons, which are counted towards the stability of both the participating atoms.

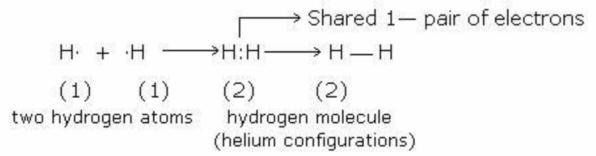
When the two atoms combine by mutual sharing of electrons, then each of the atoms acquires a stable configuration of the nearest noble gas. The compounds formed due to covalent bonding are called covalent compounds. The shared pair of electrons are called Bond Pairs.

Formation of Covalent Bonds

The Hydrogen Molecule

The hydrogen atom (Atomic number = 1) has 1 electron in the K shell. It tries to acquire the configuration of He (Atomic number 2). This is possible if the two combining atoms share their valence electron to form one covalent bond between themselves.

H - H ; H: H 1 covalent bond between two 2 H atoms forms H2



The Oxygen Molecule

Oxygen (Atomic number 8) has 6 valence electrons, 2 short of the octet configuration. The two oxygen atoms share two pairs of electrons to form 2 covalent bonds between them.

$(O = O,\ O :^{\times}_{\times} O)$

To form Oxygen molecule O_2.

Covalency

The number of electrons, which an atom contributes towards mutual sharing during the formation of a chemical bond, is called its covalency in that compound. Thus, the covalency of hydrogen in H_2 (H - H, H : H) is one; that of oxygen in O_2 is two $(O = O, O :^{\times}_{\times} O)$ and that of nitrogen in N_2 is three $(N \equiv N, N \vdots \begin{smallmatrix}\times\\ \times\\ \times\end{smallmatrix} N)$.

Sometimes one or more pairs of electrons in the valence shell of the atom do not take part in bonding, and are known as a lone pairs; they are also called non-bonding pair of electrons.

Example:

Each atom of oxygen has 2 pairs of non bonding electrons.

Multiple Covalent Bonds

The covalent bonds developed due to mutual sharing of more than one pair of electrons, are termed 'multiple covalent bonds'. These are:

Double covalent bond

The bond formed between two atoms due to the sharing of two electron-pairs is called a double covalent bond or simply a double bond. It is denoted by two small horizontal lines (=) drawn between the two atoms, e.g., O = O,

O = C = O etc.

Triple covalent bond

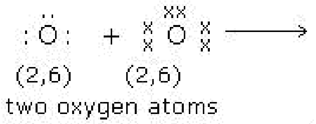
Bond formed due to the sharing of three electron pairs is called a triple covalent bond or simply a triple bond. Three small horizontal lines between the two atoms denote a triple bond e.g.,

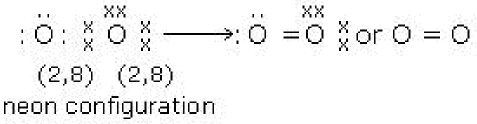
N≡N, and H-C≡C-H (acetylene).

Formation of Molecules Having Double Bond

Formation of oxygen (O_2) molecule

Each oxygen atom has six electrons in its valence shell. Thus, it requires 2 more electrons to achieve the nearest noble gas configuration. This is achieved by sharing two pairs of electrons by the two oxygen atoms as shown below:



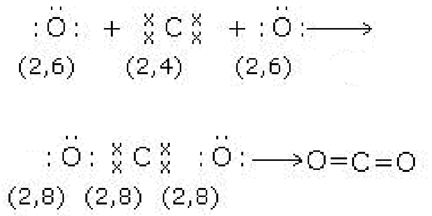


Formation of carbon dioxide molecule (CO_2)

The electronic configurations of carbon and oxygen are,

C 2, 4 and O 2, 6

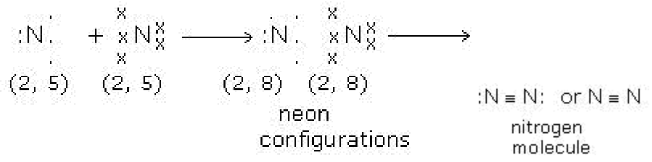
Thus, each carbon atom requires four, and each oxygen atom requires two more electrons to acquire noble gas configurations. This is achieved as follows:



Formation of molecules having triple bond

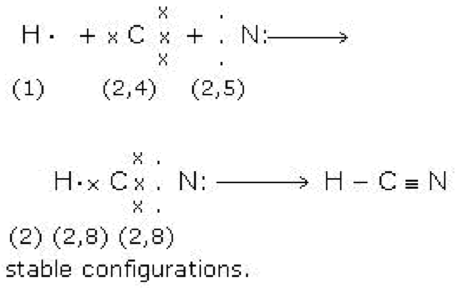
Formation of nitrogen (N_2) molecule

Nitrogen atom has five electrons in its valence shell. Thus, it requires three more electrons to acquire a stable configuration of the nearest noble gas (neon). This is done by mutually sharing three pairs of electrons as shown below:



Formation of hydrogen cyanide (HCN) molecule

The carbon atom in HCN, shares one electron-pair with hydrogen, thus forming a single covalent bond with H atom. The C atom shares three electron pairs with N atom to form a triple bond between C and N. The combining of atoms and Lewis structure of HCN molecule is given below:



General Properties of Covalent Compounds

The main characteristic properties of covalent compounds are:

- State of existence: The covalent compounds do not exist as ions but exist as molecules. There are weak intermolecular forces between the molecules and hence they exist as liquids or gases at room temperature. However, a few compounds also exist in the solid state e.g., urea, sugar etc.
- Low melting and boiling points: The melting and boiling points of covalent compounds are generally low. This is because of the fact that the forces between the molecules are weak and thus are easily overcome at low temperatures.
- Solubility: Covalent compounds are generally insoluble or less soluble in water and in other polar solvents. They are however, soluble in non-polar solvents such as benzene, carbon tetrachloride etc.
- Non-conductors: Since covalent compounds do not give ions in solution, these are poor conductors of electricity in the fused or dissolved state.
- Molecular reactions: The reactions between covalent compounds occur between their molecules. These involve the breaking of covalent bonds in reacting molecules and forming new covalent bonds to give molecules of the products. These reactions are quite slow because energy is required to break covalent bonds.

Occurrence of Metals

Minerals and Ores

Metals and their compounds are found in the earth's and are known as minerals. Ores are minerals from which metals are extracted profitably and conveniently. Ores contain metal compounds with a lower percentage of impurities. All the ores are minerals, but all minerals are not necessarily ores.

Types of ores

Oxides	Carbonates	Halides	Sulphides	Sulphates
Zincates (ZnO)	Marble or limestone ($CaCO_3$)	Fluorspar (CaF_2)	Zinc blende (ZnS) Galena (PbS)	Anglesite ($PbSO_4$)
Haematite ($Fe_2O_3.xH_2O$)Magnetite (Fe_3O_4)	Calamine ($ZnCO_3$)	Cryolite (Na_3AlF_6)	Iron pyrites (FeS_2)	Barium sulfate ($BaSO_4$)
Bauxite ($Al_2O_3.2H_2O$)	Siderite	Horn Silver	Cinnabar (HgS)	Gypsum

	($FeCO_3$)	($AgCl$)		($CaSO_4.2H_2O$)
Cuprite (Cu_2O)	Magnesite ($MgCO_3$)	Rock salt ($NaCl$)		Epsom salt ($MgSO_4.7H_2O$)

In the Free State

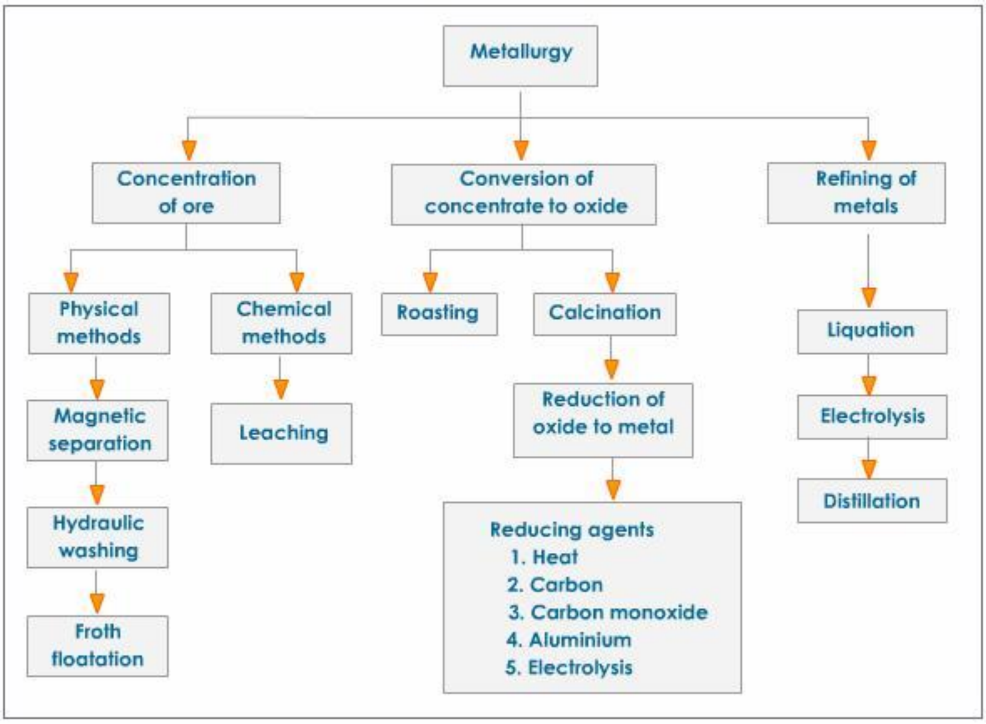
Very few metals exist in the free or native state. Only metals like gold, platinum and mercury are occasionally found in the Free State i.e., in the pure form. Sometimes, copper and silver may also be found in the Free State. Such metals are not acted upon by air or water.

In the Combined State

The rest of the metals occur in the combined form as compounds such as oxides, carbonates, sulphides, sulphates, silicates, chlorides, nitrates, phosphates etc. Copper and silver are two metals which occur in free as well as combined state as sulphide, oxide or halide ores. Metals at the top of the activity series (K, Na, Ca, Mg and Al) are so reactive that they are never found in nature as free elements. The metals in the middle of the activity series (Zn, Fe, Pb, etc.) are moderately reactive. They are found in the earth's crust mainly as oxides, sulphides or carbonates.

Extraction of Metals - Metallurgy

The various processes involved in the extraction of metals from their ores and their subsequent refining are known as metallurgy.



Concentration - Enrichment of Ores

Ore is an impure metal containing large amount of sand and rocky material. The impurities like sand, rocky materials, limestone, mica etc present in the ore is called **gangue or matrix.** These impurities must be removed from the ore before the extraction of the metal. The substance added to the ore to remove the matrix called **flux** results in the formation of a fusible compound called **slag.**

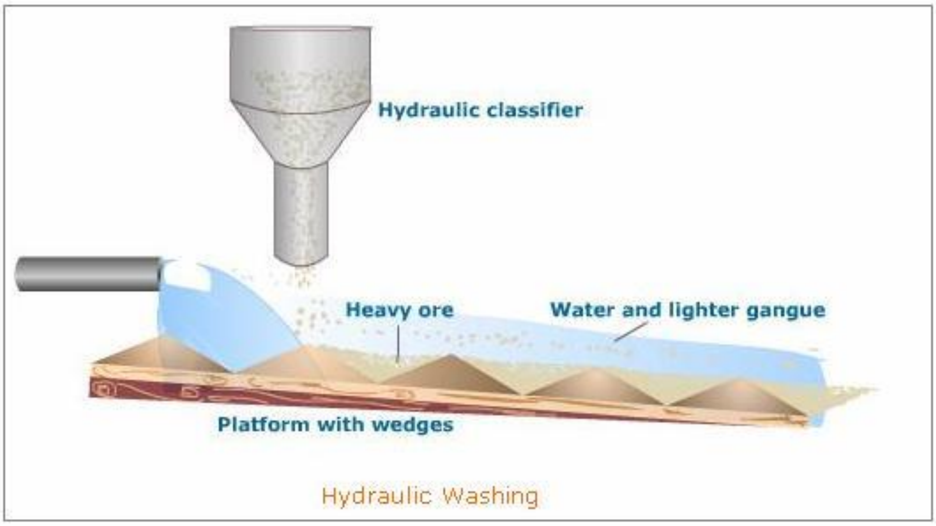
Gangue + Flux = Slag.

The processes used for removing the gangue from the ore are based on the differences between the physical or chemical properties of the gangue and the ore. At first the ore is crushed to powder. The **pulverized** ore is separated by **physical processes** like hydraulic washing, froth-floatation, and magnetic separation or by **chemical processes**, depending on the nature of the ore and its impurities. Concentration of the ore is also known as 'dressing' or 'enrichment' of ore.

Physical Methods of Concentration

Hydraulic Washing (Gravity Separation)

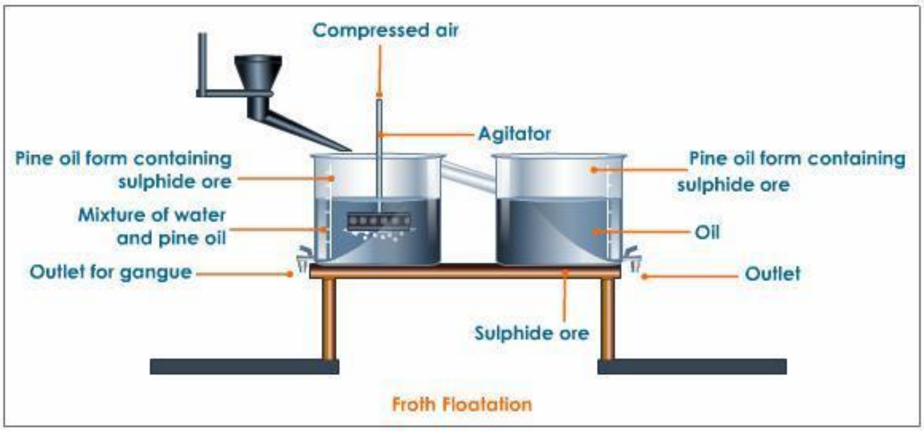
The ore particles are poured over a hydraulic classifier which is a vibrating sloped table with grooves and a jet of water is allowed to flow over it. The denser ore settle in the grooves while the lighter gangue particles are washed away. This method is used for concentration of heavy oxide ores of lead, tin, iron etc.



Hydraulic Washing

Froth Floatation

This method employs a mixture of water and pine oil which is made to froth in a tank to separate sulphide ores. The differences in the wetting properties of the ore and gangue particles separate them.

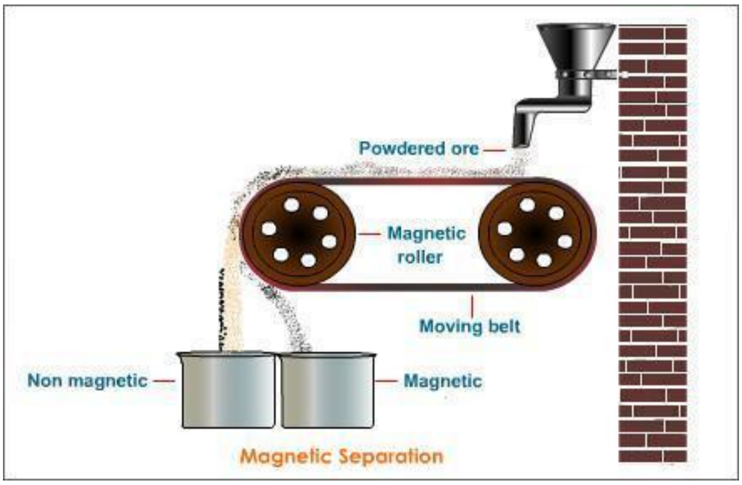


Froth Floatation

A mixture of water, pine oil, detergent and powdered ore is first taken in a tank. A blast of compressed air is blown through the pipe of a rotating agitator to produce froth. The sulphide ore particles are wetted and coated by pine oil and rise up along with the froth (froth being lighter). The gangue particles wetted by water sink to the bottom of the tank (water being heavier). Sulphide being more electronegative attracts the covalent oil molecules. The gangue being less electronegative is attracted by the water. The froth containing the sulphide ore is transferred to another container, washed, and dried.

Magnetic Separation

Magnetic ores like pyrolusite (MnO_2) and chromite ($FeO.Cr_2O_3$) are enriched by this method by making use of the difference in the magnetic properties of the ore and gangue particles.



The powdered ore is dropped over a conveyor belt running over two rollers, one of which is magnetic. The magnetic ore particles get attracted to the magnetic roller and run along with the conveyor belt for a little longer than the non magnetic gangue. Gangue particles drop down first forming a heap. Then, the magnetic ore particles drop down forming a separate heap. Thus, two separate heaps of ore and gangue particles are formed.

Extracting Metals Low in the Activity Series

Metals that are low in the activity series are very un-reactive. The oxides of these metals can be reduced to metals by heating alone. For example, mercury is obtained from its ore, cinnabar (HgS), by the process of heating.

Reduction Using Heat

$$2HgS + 3O_2 \xrightarrow{Heat} 2HgO + 2SO_2$$

$$2HgO \xrightarrow{Heat} 2Hg + O_2$$

Mercuricoxide

Copper can also be obtained in a similar manner from its sulphide ore (Cu_2S).

Note: The oxides of Mercury & copper metals decompose on heating.

Extracting Metals in the Middle of the Activity Series

Metals such as iron, zinc, lead, copper, etc., are in the middle of the reactivity series. These are moderately reactive metals and are usually present as sulphides or carbonates.

These metals are obtained from their ores by the processes of reduction. In the reduction process, it is the oxide ore that is reduced.

It is easier to reduce an oxide ore as compared to its sulphides and carbonates. If the ore is not an oxide ore, it is first converted to the oxide by the process of calcination or by roasting.

Roasting

Sulphide ores are converted into oxides by heating strongly in the presence of excess air, so that oxygen gets added to form the corresponding oxides. Sulphur impurities escape as gas. This process is known as roasting.

$$2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$$
$$2PbS(s) + 3O_2(g) \rightarrow 2PbO(s) + 2SO_2(g)$$
$$4FeS(s) + 7O_2(g) \rightarrow 2Fe_2O_3(s) + 4SO_2(g)$$

Calcination

In this process the ore is heated to a high temperature in the absence of air, or where air does not take part in the reaction. Usually, carbonate ores or ores containing water are calcined to drive out carbonate and moisture impurities.

$$ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$$
$$Al_2O_3.2H_2O(s) \rightarrow Al_2O_3(s) + 2H_2O(l)$$

Differences between Roasting and Calcination

Roasting	**Calcination**
Heating in the presence of air	Heating in the absence of air
Sulphide ores are roasted Sulphur dioxide is released $2ZnS + 3O_2 \xrightarrow{\Delta} 2ZnO + 2SO_2\uparrow$	Carbonate ores are calcined. Carbon dioxide is released $CaCO_3 \xrightarrow{\Delta} CaO + CO_2\uparrow$
Is done at high temperatures, sometimes higher that the melting point of the ore.	Is done at lower temperatures, generally below the melting point of the ore.
Purpose is to remove impurities as volatile oxides & to oxidize the ore and to remove any moisture that may be present.	Purpose is to remove impurities as volatile oxides and to remove the moisture present.

Reduction of Ores

The oxide obtained by calcination or roasting is then reduced by either carbon or hydrogen. Carbon is usually used in the form of coke, or carbon monoxide. Not all metallic oxides can be reduced by carbon or carbon monoxide. Thus, in order to extract metal from its ore the method adopted depends on the reactivity of the metal i.e., its position in the metal activity series.

Metal	Method of reduction of oxide
Zn	By carbon reduction method. This metal can be reduced by coke only
Fe, Pb, Cu	By carbon reduction method. These metals can be reduced by coke as well as carbon monoxide. They can also be reduced by hydrogen.

Reduction Using Carbon (Coke)

This method is used for oxides of moderately reactive metals. Coke is an inexpensive reducing agent and is most widely used.

$$2ZnO + C \xrightarrow{\Delta} 2Zn + CO_2\uparrow$$
$$2PbO + C \xrightarrow{\Delta} 2Pb + CO_2\uparrow$$

Reduction Using Carbon Monoxide

Carbon monoxide is a powerful reducing agent. It is used for the reduction of hematite in the blast furnace.

$$Fe_2O_3 + 3CO \xrightarrow{\Delta} 2Fe + 3CO_2\uparrow$$

Extracting Metals towards the Top of the Activity Series

Metals such as sodium, magnesium, calcium, aluminum high up in the reactivity series are very reactive and cannot be obtained from their compounds by heating with carbon. This is because these metals have more affinity for oxygen than carbon. These metals are obtained by electrolytic reduction.

Reduction by Electrolysis

This process is used for oxides of highly reactive metals that are above aluminum in the reactivity series. It is also used for reduction of aluminum oxide. Example:

$$2Al_2O_3 \xrightarrow{\text{electricity}} 4Al + 3O_2$$

Alumina has a very high melting point and the cost of maintaining the electrolyte in the molten state is very high. However, if alumina is mixed with cryolite (Na_3AlF_6) and fluorspar (CaF_2) the melting point is lowered drastically and the cost reduced too. These substances increase the conductivity of the electrolyte.

Reduction Using Aluminum (Aluminothermy)

Aluminum is an expensive reducing agent. It is used for reduction of oxides of highly reactive metals. But these metals need to be below aluminum in the reactivity series. Example:

$Cr_2O_3 + 2Al \xrightarrow{\Delta} Al_2O_3 + 2Cr$

Aluminothermy is also used for welding broken iron rails. As the reaction is highly exothermic, the heat generated in the reaction causes the metallic iron formed to melt. This molten iron drips down over the two pieces to be welded and joins them on solidifying.

$Fe_2O_3 + 2Al \xrightarrow{\Delta} Al_2O_3 + 2Fe$

Refining of Metals

Most metals obtained by the reduction process are not very pure. These have to be further refined or purified. Purification of the metal is the last step in metallurgy. Refining is based on the difference between the properties of metals and their impurities. The following process we use for refining.

Electro refining

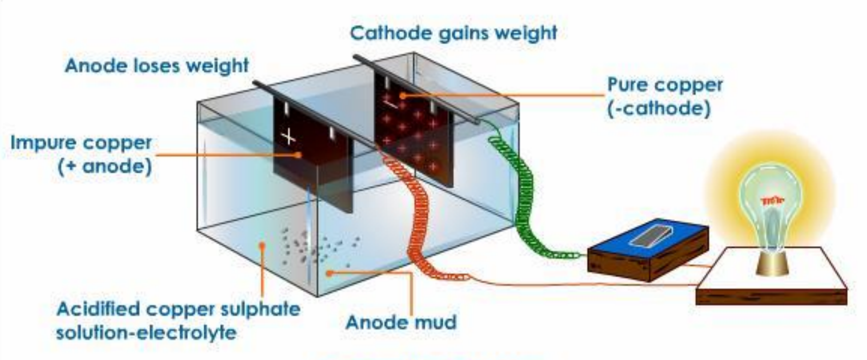
Electrolysis can be used for both extractions of metal (which cannot be separated by chemical reduction process) as well as for further purification of metals obtained by any other method. In the electro refining process a block of impure metal is made the **anode** and a thin sheet of pure metal is made the **cathode.** The electrolytic cell contains an aqueous solution of the metal salt. When electric current of a suitable voltage is passed, impure metal at the anode gets dissolved to deposit the pure metal at the cathode. Metal ions from the anode enter the electrolyte as follows:

$M \rightarrow M^{+n} + ne^-$

These ions get deposited on the cathode as follows

$M^{+n} + ne^- \rightarrow M$

The impurities are left behind as anode mud near the anode. The anode finally disintegrates while the cathode gains in weight due to the collection of pure metal.



Electrorefining of Metals

This method is used for refining volatile metals like copper, silver, tin, nickel that have boiling points lower than their impurities. e.g., zinc, mercury.

- An electrolyte is a compound (salt, acid or base), which in solution or in a molten state conducts an electric current and is simultaneously decomposed by it.

Electrolytes are ionized into electrically charged ions, which carry the current.

- Charged ions move towards the oppositely charged electrodes to give up their electric charge and become atoms; these are either liberated or deposited at the electrodes.

Corrosion of Metals

We have learnt that metals that are chemically active get corroded in the presence of a moist atmosphere. Corrosion is an oxidation reaction with atmospheric oxygen in the presence of water on the surface of a metal. Iron corrodes more quickly than most other transition metals to form an iron oxide. Corrosion or rusting of iron is accelerated in the presence of CO_2 and also in the presence of salt solution.

Rusting is $Fe_{(s)} + O_{2(g)} + H_2O_{(l)} \Rightarrow Fe_2O_3.xH_2O_{(s)}$

i.e., rust is hydrated iron (III) oxide

Rusting is oxidation because it involves iron gaining oxygen ($Fe \Rightarrow Fe_2O_3$) or iron atoms losing electrons ($Fe - 3e^- \Rightarrow Fe^{3+}$). The equation is not meant to be balanced and the amount of water 'x' is variable, from dry to soggy.

The major problem of corrosion occurs with iron (or steel) as it is used as a structural material in industries like construction, infrastructure, bridges, rail transport power transmission, shipbuilding, automobiles, heavy industries etc.

Aluminum, another useful structural metal, also undergoes an oxidation reaction, but does not oxidize and corrode as quickly as its reactivity suggests. Once a thin oxide layer of Al_2O_3 has formed on the surface, it forms a barrier to oxygen and water to prevent further corrosion of the aluminum. Hence aluminum is called a **self-protective metal**. Aluminum can be made harder, stronger and stiffer by mixing it with small amounts of other metals (e.g., magnesium) to make alloys.

The alkali metals like sodium used in chemistry laboratories and in some chemical industries rapidly corrodes in air and need to be stored under oil.

Copper and lead are both used in roofing situations because neither is very reactive. The compounds formed on the surface do not flake away as easily as rust does from iron. Lead corrodes to a white lead oxide or carbonate and copper corrodes to form a basic green carbonate (combination of the hydroxide $Cu(OH)_2$ and carbonate $CuCO_3$). In the past both metals have been used for piping but as lead is considered too toxic copper is usually used.

Non-reactive metals like gold, platinum, mercury do not corrode.

Prevention of Corrosion

Iron and steel (alloy of iron) are most easily protected by paint which provides a barrier between the metal and air/water. Moving parts on machines can be protected by a water repellent oil or grease layer. Covering the surface with enamel and lacquers is another method.

Sacrificial Protection

'Rusting' can be prevented by connecting iron to a more reactive metal (e.g., zinc or magnesium). This is referred to as sacrificial protection or sacrificial corrosion, because the more reactive protecting metal is preferentially oxidized away, leaving the protected metal intact.

Alloying

Iron or steel along with other metals can also be protected by 'alloying' or mixing with other metals (e.g., chromium) to make non-rusting alloys. Stainless steel is an example of a non-rusting alloy of iron and carbon. Brass, an alloy containing copper is another metal alloy which is less expensive and non-reactive.

Galvanizing

Coating iron or steel with a thin zinc layer is called 'galvanizing'. This layer is produced by electrolytic deposition. Dipping the iron/steel object in molten zinc and using it as the negative cathode zinc is coated on it. Zinc preferentially corrodes or oxidizes to form a zinc oxide layer that does not flake off like iron oxide rust. Also, if the surface is scratched, the exposed zinc again corrodes before the iron and continues to protect it.

Electroplating

Coating the surface with metals like tin, chromium, nickel etc. by electroplating is also utilized to prevent corrosion. Steel cans are protected by relatively un-reacted tin and works well as long as the thin tin layer is complete.