Metals and non-metals

Metal

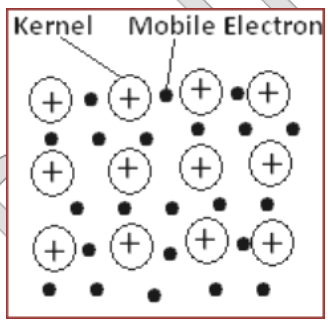
Metals are crystalline solids and atoms are linked to one another by special type of bonds known as metallic bonds. "***Metals are those elements which can release and form cation.***"

Metallic bond: It is the bond between atoms of a metal.

A metal atom may be supposed atom be made up of two parts:

1. ***Valence electrons*** – Electron present in valence shell are known as ***valence electron***. These can be easily removed from atom.
2. ***Kernel*** – kernel include nucleus and all other electron shells except valence shell in an atom. The kernel is positively charged.

*A metallic object may be regarded as the sea of valence electron in which positively charged **kernels** are immersed. Each kernel is surrounded by a number of **Valence electron** and vice versa. **Valence electron** is mobile and can move from one position to another. Kernel has negligible movement. The force of attraction between kernel and **Valence electron** leads to formation of **metallic bonds**.*

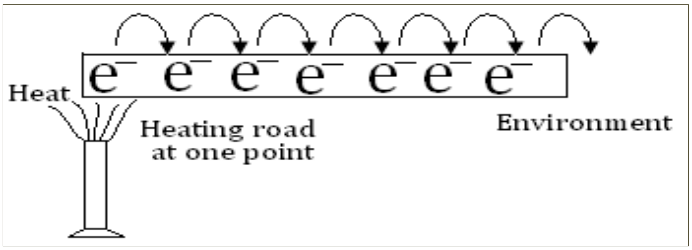


Physical properties of metal:

1. ***Metallic luster***: In their pure state, metals have a lustrous (shiny) surface. This property is called ***metallic luster***.

   When light falls on the surface of a metal, the atoms absorb photons as energy. They get exited and start vibrating. These is vibrating electrons release energy as light, therefore, metal surface shines.
2. ***Hardness:*** Metals are usually very ***hard***. The hardness of the metals depends upon the strength of metallic bond. Stronger the metallic bond, harder is the metal.

   Exception: **Na**, **Mg**, **K** are so soft they can be cut by a knife. **Hg** is liquid at room temperature.
3. ***Malleability:*** Some metals can be beaten into thin sheets, this property is called malleability. **Au** and **Ag** are most malleability.
4. ***Thermal conductivity:*** metals are good conductor of heat due to presence of free electrons. The metals take heat or current and pass it on to another electron. This passing goes on while the last electron gets heat.



5. ***Ductility:*** The property due to which a metal can be drawn into thin wires is known as ductility. **Cr**, **Al**, **Mg** and **Fe** are most ductile.

6. ***Electric conductivity:*** metals are good conductor of electricity. They differ in their conducting power. It is due to presence of free Electrons which are ***carrier of charge***.
7. ***State:*** metals are solids at room temperature only **Hg** is liquid at room temperature.
8. ***Density:*** metals have high densities except **Na** & **K**.
9. ***Sonorous:*** metals are sonorous means they have ringing sound. Metals are used for making bell's and musical instrument.

Non metals

"Non metals are those elements which form anion by accepting electrons." The most abundant non metal is on the earth is ***oxygen*** which constitutes about ***50%*** of earth crust whereas the ***silicon*** is the ***second most abundant metal*** constituting ***26%*** of the earth crust.

Physical properties of Non-metals

1. ***Malleability and ductile:*** Non metals don not have malleability and ductility.
2. ***Brittleness:*** Non metals are hammered, they break into pieces.
   Diamond (allotrope of C) is of ductile nature but it is quite hard
3. ***Luster:*** metals do not have luster because their atoms have loosely attached electron.
   Diamond, iodine, graphite are lustrous
4. ***Hardness:*** non metals are not hard
   Sulphure, phosphorous and **graphite** are soft.
   Diamond is very had (hardest substance on earth)
5. ***Electric conductivity:*** non metals are poor conductors of heat and electricity because they do not have free electrons to conduct electricity.
   Graphite is good conductors of electricity because the carbon atoms in graphite have one free electron. Each not involved in the bond formation
6. ***Density:*** Non-metals have low density because atoms are linked by covalent bonds, and not closely packed.
7. ***Melting and boiling point:*** Non metals have low melting and boiling point.
8. ***State:*** They are found in all-three state
   - **H_2, O_2, N_2, Ne** and **Ar** are gaseous at room temperature.
   - **Boron, Si** and **C** are solid at room temperature.

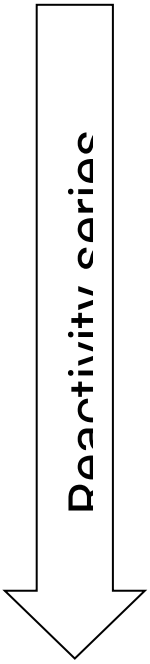
Exceptions

1. Mercury (metals) is liquid at room temperature.
2. Crystals (non-metals) have bright luster (non-metal)
3. Graphite (non-metal) is good conductor of electricity.
4. Li, Na and K (metals) are very soft
5. Diamond (non-metal) is hardest substance.
6. Gallium and calcium (metals) have very low melting point. They will melt on our palm.
7. Diamond (non – metal) have melting point.
8. Li, K, Na have low densities, low melting point.

Chemical properties of metals

The chemical properties of the metals are depend upon their tendency of release electron. Greater the tendency more will be the reactivity of the metal.

Reactivity series	
K	Potassium
Na	Sodium
Ba	Barium
Ca	calcium
Mg	Magnesium
Al	Aluminium
Zn	Zinc
Cr	Chromium
Fe	Iron
Cd	Cadmium
Co	Cobalt
Ni	Nickel
Sn	Tin
Pb	Lead
H	Hydrogen
Cu	Copper
Hg	Mercury
Ag	Silver
Au	Gold
Pt	platinum



1. **Active metals:**
   Metal that are place above Hydrogen evolve the gas on reacting with dilute acid, they are know as ***active metals.***
2. **Inactive metals:**
   The metals that are place below hydrogen do not evolve gas on reacting with dilute acid. They are known as ***inactive metals***.
3. **Nobel metals:**
   Pt and **Au** are very little reactive. Although **H** is a non metal it has been included in the reactivity series it behaves like metals and forms H^+ by loosing its only electron.

Why Some Metals are More Reactive and Other Less reactive?

When metals react, they lose electrons to form positive ions. Now, if a metal atom can lose electron easily to form positive ions, it will react rapidly with other substances and hence it will be a reactive metal. On the other hand, if a metal atom loses electrons less readily to form positive ions; it will react slowly with other substances. Such a metal will be less reactive.

Example: **Na** atoms lose electrons readily to form sodium ions, due to which sodium metal is very reactive. On the other hand, **iron** atoms lose electrons less readily to form positive ions, so iron metal is less reactive.

Metals which are More Reactive than Hydrogen

Those metals which lose electrons **more readily** than ***Hydrogen*** are said to be more reactive than Hydrogen. All the metals which have been placed above hydrogen in the activity series lose electrons more readily than hydrogen, and hence they are more reactive than Hydrogen.

Metals which are Less Reactive than Hydrogen

Those metals which lose electrons **less readily** than Hydrogen are said to be less reactive than hydrogen. All the metal placed below hydrogen in the reactivity series loses electrons less readily than hydrogen, and hence they are less reactive than hydrogen.

A. Reaction of metals with oxygen:

When metals are burnt in air they react with O_2 of air to from ***metal oxide***.

Metal + oxygen ⟶ ***metal oxide.***

Metal oxides are ***basic*** in nature. Because they dissolve in H_2O to from bases, these are called ***alkalis***.

Example

1. **K** and **Na** react with O_2 at room temperature.

   $4Na + O_2 \longrightarrow 2Na_2O$

   $4K + O_2 \longrightarrow 2K_2O$

   Both these metals are so reactive that they react ***violently***. As a result heat evolved, they ***catch fire***.

2. **Mg** reacts with O_2 upon heating and burns brightly

   $2Mg + O_2 \longrightarrow 2MgO$

3. **Fe** reacts with O_2 but do not burn. The product is a mixture of **FeO. Fe_2O_3** *(Iron. II, III oxide)* or **Fe_3O_4** *(ferroso ---ferric—oxide).* It is brown in colour.

   $3Fe + 2O_2 \longrightarrow$ FeO. Fe_2O_3 or Fe_3O_4 *(brown)*

4. **Cu** reacts with O_2 on prolonged heating to form cupric oxide (CuO)

   $2Cu + O_2 \longrightarrow 2CuO$ *(Cupric Oxide- Black coloured)*

5. **Ag, Au** do not react with O_2 even at very high temperature.
6. **Al** reacts with O_2 to **Al_2O_3** *(Aluminum Oxide)*

   $4Al + O_2 \longrightarrow 2Al_2O_3$

~: Nature of Metallic-Oxide: ~

Oxide of metals dissolve in water to from soluble hydroxides known s ***alkalies.***

1. $Na_2O + H_2O \longrightarrow 2NaOH$ (aq)
2. $K_2O + H_2O \longrightarrow 2KOH$ (aq)
3. $MgO + H_2O \longrightarrow Mg(OH)_2$ (aq)
4. $CaO + H_2O \longrightarrow Ca(OH)_2$ (s)
5. $Al_2O_3 + 3H_2O \longrightarrow 2Al(OH)_3$ (s)

$Ca(OH)_2$ and $Al(OH)_3$ *are insoluble in water and known as* ***bases.***

Amphoteric Oxides

The oxides which have the characteristic of both acids and bases are known as amphoteric oxide

Example: Al_2O_3, ZnO

1. ***Behaving as base***

   $Al_2O_3 + 6HCl \longrightarrow 2AlCl_3 + 3H_2O$

   $ZnO + 2HCl \longrightarrow ZnCl_2 + H_2O$

2. ***Behaving as acid***

   $Al_2O_3 + 2NaOH \longrightarrow 2NaAlO_2 + H_2O$

   (Sodium-aluminate)

   $ZnO + 2NaOH \longrightarrow Na_2ZnO_2 + H_2O$

   (Sodium-Zincate)

B. Reaction of metals with water

Metals reacts with water and produce metal-hydroxide and hydrogen gas

Metal + Water ⟶ ***Metal Hydroxide + Hydrogen gas***

When a metal reacts with steam, it forms metal oxide and hydrogen gas

Metal + steam ⟶ ***metal oxide + hydrogen***

Actually whenever metal react with water first it forms metallic oxide which is soluble in water and form metal hydroxide. Hydrogen is liberated in the first step of reaction.

Metal+ water ⟶ *metal-oxide + Hydrogen*

Metal oxide + water ⟶ *metal hydroxide*

I. $Ca + H_2O \longrightarrow CaO + H_2$

II. $CaO + H_2O \longrightarrow Ca(OH)_2$

In combined form:

$Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2$

1. **Na** and **K** react with *water* violently even in the cold and immediately ***catch fire***. So they are kept immersed in **kerosene oil** which act as barrier between humidity of atmosphere and prevent fire.
When **Na** reacts with water ***Hydrogen gas*** evolved. This reaction is ***exothermic***. The ***ignition*** temperature of hydrogen is very ***low***, therefore, it is highly combustible in nature and catches fire.

$2K + 2H_2O \longrightarrow 2KOH\ (aq) + H_2 + \Delta$

$2Na + 2H_2O \longrightarrow 2\ NaOH\ (aq) + H_2 + \Delta$

2. **Calcium** reacts with water less violently.

$Ca\ (s) + 2H_2O\ (l) \longrightarrow Ca(OH)_2\ (aq) + H_2$

3. **Mg** reacts with water upon heating.

$Mg\ (s) + 2H_2O\ (g) \longrightarrow Mg(OH)_2\ (aq) + H_2$

The densities **Mg** and **Ca** are greater than water. However, both float over water. Because bubble of hydrogen gas evolved during the reaction stick on the surface of metals and they starts floating.

4. Metals like **Al**, **Fe** and **Zn** do not react with either cold or hot water but they react with **steam** to form metal oxide and hydrogen.

$2Al + 3\ H_2O\ (g) \longrightarrow Al_2O_3 + 3H_2$

$3\ Fe + 4\ H_2O\ (g) \longrightarrow Fe_3O_4 + 4H_2$

5. Metals like **Cu**, **Ag**, **Au**, **Pt** and **Pb** do not react with water under any condition.

C. Reaction of metals with acid:

1. Metals place above Hydrogen in the reactivity series react with dilute acids like **HCl**, **H_2SO_4** to evolve H_2 gas and forming ***metal salts***.

Metal + dil. Acid ⟶ ***Metal salt + Hydrogen gas***

Example:

I. $Ca + 2HCl \longrightarrow CaCl_2 + H_2$

II. $Mg + H_2SO_4 \longrightarrow Mg\ SO_4 + H_2$

III. $2Na + 2HCl \longrightarrow 2NaCl + H_2$

The electron removed by metal are accepted by ***H^+*** *ions of acid to evolve* **H_2** *gas.*

$Ca \longrightarrow Ca^{2+} + 2e^-$

$2H^+\ (aq) + 2Cl^-\ (aq) + 2e^- \longrightarrow H_2 + 2Cl^-$

2. When metals reacts with Nitric acid **(HNO_3)** they form **H_2O** instead of **H_2 gas** because **HNO_3** is a strong oxidizing agent which oxidizes the **H_2** gas to form **water** and it self reduced to any of its **oxides(** Nitrous oxide**(N_2O),**Nitric oxide**(NO),**Nitrogen oxide**(NO_2)).**

I. $Al + 6HNO_3 \longrightarrow Al(NO_3)_3 + 3H_2O + 3NO_2$

II. $Zn + 4HNO_3 \longrightarrow Zn(NO_3)_2 + 2H_2O + 2NO_2$

a. $Zn + 2HNO_3 \longrightarrow Zn(NO_3)_2 + 2H$

b. $[\ H + HNO_3 \longrightarrow NO_2 + H_2O\ ] \times 2$

3. Mn, Mg do react with very dilute HNO_3 to evolve H_2 gas.

$$Mg + 2HNO_3\ (very\ dil.) \longrightarrow Mg(NO_3)_2 + H_2$$

Aqua-Regia (royal water)

It is the mixture of concentrated **HCl** and concentrated **HNO_3** in the ratio of **3:1** by volume. They combined to evolve ***nascent chlorine*** which reacts with both gold, platinum reacts with both gold, platinum. As a result, the dull layer from the surface of the metal is removed and it again acquires a shining book.

$$\underbrace{3HCl + HNO_3}_{(Aqua\ Regia)} \longrightarrow \underset{(Nitrosyl\ chloride)}{NOCl} + 2H_2O + Cl\ [Nascent]$$

$$Au + 3Cl \longrightarrow AuCl_3 \quad (Auric\text{-}chloride)$$

$$Pt + 4Cl \longrightarrow PtCl_4 \quad (Platinic\ chloride)$$

*The word **'nascent'** means just born in these reactions chlorine which is evolved in the reaction is not separated and reacts as such in the mixture. That is why it is known as nascent chorine. It is more reactive then its molecular form.*

D. Reaction of metals with solution of other Salts:

A metal placed higher in the activity series can displace the metal which occupies a lower position from the aqueous solution of its salt.

1. **Zinc** displaces **copper** from **$CuSO_4$** solution.

$$\underset{(Bluish\text{-}Colour)}{Zn\ (s) + CuSO_4\ (aq)} \longrightarrow \underset{(colour\ less)}{ZnSO_4\ (aq) + Cu\ (s)}$$

$$Zn\ (s) \longrightarrow Zn^{+2}\ (aq) + 2e^-$$
$$Cu^{2+}\ (aq) + 2e^- \longrightarrow Cu\ (s)$$

2. **Iron** displaces **Copper** from **$CuSO_4$.**

$$Fe + CuSO_4 \longrightarrow \underset{(Light\ green)}{FeSO_4 + Cu}$$

3. **Zinc** displaces **Fe** from **iron sulphate.**

$$Zn + FeSO_4 \longrightarrow ZnSO_4 + Fe\ (s)$$

4. Cu displace Ag from $AgNO_3$.

$$Cu + 2AgNO_3\ (aq) \longrightarrow Cu(NO_3)_2\ (aq) + 2Ag\ (s)$$

Chemical Properties of Non-metals

The chemical properties of non metals are depend upon the electronic configuration of their atoms.
Non metals from negative ions or anion and called electronegative element.

$$N + 3e^- \longrightarrow N^{3-}$$
$$O + 2e^- \longrightarrow O^{2-}$$
$$C + 4e^- \longrightarrow C^{4-}$$

The chemical prosperities of metals and non metals are different because of electronic configuration of their atoms. Whereas metals atoms tend to lose one or more valence electrons, the atoms of non metals tend to accept these.

A. Reaction with Oxygen of Non Metals:

Non- metals react with air or oxygen upon hating to form oxides.

1. $C\ (s) + O_2\ (g) \longrightarrow CO_2\ (g)$

2. $S_8\ (s) + 8O_2\ (g) \longrightarrow 8SO_2\ (g)$

If SO_2 is hated in the presence of vanadium (V) oxide or ***vanadium pentoxide*** (V_2O_5) catalyst to 273 k. SO_2 oxidized to sulphure tri-oxide.

$$2SO_2 + O_2 \longrightarrow 2SO_3(g)$$

Nature of the Non-metalic oxides

Oxides of non-metals are ***acidic*** in nature because they form acid when dissolve in water.

1. $CO_2 + H_2O \longrightarrow H_2CO_3$ *(carbonic acid)*
2. $SO_2 + H_2O \longrightarrow H_2SO_3$ *(sulphurous acid)*
3. $P_2O_{10} + 6H_2O \longrightarrow 4H_3PO_4$ *(phosphoric acid)*

S *and* ***P*** *exit in molecular form as an aggregate of eight and four atoms respectively. They are not present in atom. In chemical reactions, both sulphure and phosphorus may be shown as atom not as molecules.*

B. Reaction with acids of Non-Metals:

Non-metals are oxidized upon heating with concentrated sulphuric acid and Nitric-acid.

Example:

$$S + 2H_2SO_4\ (conc.) \xrightarrow{Heat} 3SO_2 + 2H_2O$$

$$S + 6HNO_3\ (conc.) \xrightarrow{Heat} H_2SO_4 + 6NO_2\ (g) + H_2O$$

(Reddish brown vapour)

$$2P + 5H_2SO_4\ (conc.) \xrightarrow{Heat} 2H_3PO_4(aq) + 5SO_2\ (g) + 2H_2O$$

(Phosphoric acid)

$$P + 5HNO_3\ (conc.) \xrightarrow{Heat} H_3PO_4\ (aq) + 5NO_2\ (aq) + H_2O$$

C. Reaction with Chlorine

Non-metals react with Chlorine upon heating to form chlorides

Example: $P + 5Cl_2 \xrightarrow{Heat} 2PCl_5$

$2S + Cl_2 \xrightarrow{Heat} S_2Cl_2$ *(Sulphur mono chloride)*

Sulphure mono chloride should have the formula SCl. However it exist a dimar with formula S_2Cl_2.

D. Reaction of Non metals with water

Non metals do not react with water to evolve hydrogen.

E. Reaction of non metals with salt solution

When non metals react with salt solution the less reactive non metal get displaced by a more reactive non metal.

$$2NaBr + Cl_2 \longrightarrow 2NaCl + Br_2\ (aq)$$

Metals react with Non metals

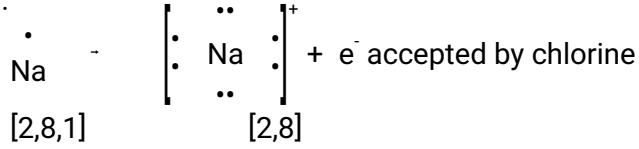
When metals react with non metals, they from ***ionic compounds*** and vice versa. From the ***electronic configurations*** of metals and non metals, it is clear that metal atoms have surplus electron in their valence shell and non metal atoms need some electrons in their valence shall to complete their outermost shell.

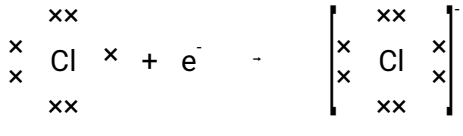
Metal atoms release electrons and form ***cation****, non metals accept electrons to form* ***anion****.*

Because of carrying opposite charges cation and anion attract each other and form bond

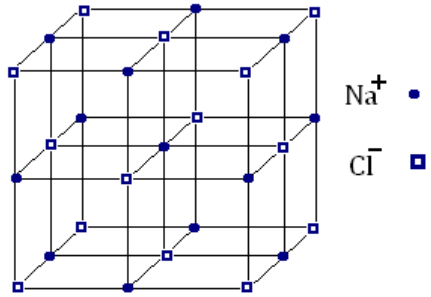
between them. The attraction force between anions and cations is known as ***electrostatic force*** which results in bond also called ***electrovalent bond***.

The compound which contains ionic bonds is their atoms are known as ionic or electrovalent compound.





$Na^+ + Cl^-$ - NaCl



Properties of ionic compounds

Ionic compounds are made of oppositely charged ions i.e. cation and anions which have strong. Electrostatic forces of attraction.

1. ***Physical state:***

Ionic compounds are generally solid and form crystals. In these crystals, ions have strong attractive forces. Crystals of different ionic compounds are different because of different arrangement of the ions known as crystal lattice.

Crystal lattices: *The regular three-dimensional arrangement of the ions in space.*

2. ***Melting and boiling points:***

Melting and boiling points are depends upon the nature of bond. Since ionic compounds have strong force of attraction they have very high melting and boiling point.

Melting and boiling points of some crystalline ionic solids

Compound	Chemical Formula	Melting Point(K)	Boiling Point(K)
Lithium chloride	LiCl	878	1570
Sodium chloride	NaCl	1074	1738
Sodium bromide	NaBr	1031	1663
Potassium Bromide	KBr	1007	1708
Potassium iodide	KI	953	1600
Calcium chloride	$CaCl_2$	1055	1870
Magnesium chloride	$MgCl_2$	987	1685
Calcium oxide	CaO	2845	3123
Magnesium oxide	MgO	3073	3873

3. ***Solubility:***
Ionic compounds are mostly water soluble. Water helps in dissociation of ions of ionic compounds.
Example: *When* ***sodium chloride*** is added to water, each ***Na^+ ion*** and ***Cl^- ion*** is surrounded by a number of ***water*** molecules. Water molecules forms envelop around these ions. These are known as ***hydrated ions***. Energy is released during hydration which helps in the separation of ions which helps in separation of Na^+ and Cl^- ions.

$$Na^+Cl^-(s) \quad - \quad Na^+(aq) + Cl^-(aq)$$

Hydrated ion

4. ***Electric conductivity:***
Ionic compounds are good conductors of electricity. Conductivity is due to the presence of ions. Ionic mobility is also required ions can make only in the molten state or in solution.
Ionic compounds do not conduct electricity in the solid state since ions do not move as they are very closely packed.

Occurrence of metals

Relative presence of metals:
Al – 8.3 , Fe – 5.1, Ca- 3.6, Na – 2.8, K – 2.6, Mg – 2.1, Others – 0.5

- Aluminum is the most abundant in earth's crust.
- Earth is the major source of metals.
- Metals can be obtained from other sources.
- Magnesium is present in chlorophyll.
- Iron is constituents of hermits.
- Sea is also sources of metals like Na, K, Ca, and Mg.
- The metals placed at the bottom of the series i.e. Ag, Au, pt generally exist in the free stat. *(* ***native state***)
- The metals at the tap of the series K, Na, Ca, Mg, Al, Fe are exist in combination of other elements *(*generally non-metals) as oxides, carbonates, sulphate, sulphides.
- The combine state of metals is known as ***minerals.*** Metal can have more than one mineral.

Ore: The minerals from where metals can be conveniently and profitably extracted are called ***ore.***
Example: the minerals of Al are:
Bauxite *($Al_2O_3.\ 2H_2O$)*, Diaspore *($Al_2O_3.\ H_2O$)*, Crinoline *(Na_3AlF_6)*, Beryl ($Be_3Al_2Si_6O_{18}$).

But the metal is generally extracted from bauxite. So bauxite is an ore of Al.

All ores are minerals whereas all mineral are not ores.

Metal	Name of Ore	Chemical formula	Chemical name
Aluminum (Al)	Bauxite	$Al_2O_3.\ 2H_2O$	Aluminum oxide
Calcium (Ca)	Dolomite	$CaCO_3MgCO_3$	Calcium-magnesium carbonate
Copper (Cu)	Malachite	$CuCO_3.Cu(OH)_2$	Basic copper carbonate
Lead (Pb)	Galena	PbS	Lead sulphide

Iron (Fe)	Haematite	Fe_2O_3	Iron (III)oxide
Silver (Ag)	Argentite	Ag_2S	Silver (I) sulphide
Sodium (Na)	Rock-salt	$NaCl$	Sodium chloride
Tin (Sn)	Cassiterite	SnO_2	Tin (IV) oxide
Zinc (Zn)	Zinc blende	ZnS	Zinc sulphide

Metallurgy: The process of extraction of metals from there ore is known as ***metallurgy.***

The actual techniques which are to be applied for the extraction of metals are linked with their position in the reactivity series.

On the basis of reactivity, metals are grouped in to three categories:

1. Metals of low reactivity
2. Metals of medium reactivity
3. Metals of high reactivity

Different techniques are to be used for obtaining the metals falling in each category.

Concentration of ores (enrichment of ores)

Ores mined from the earth are usually concentrated with large amount of impurities (soil, sound) these impurities are known as ***gangue.*** This process is known as ***enrichment of ores***.

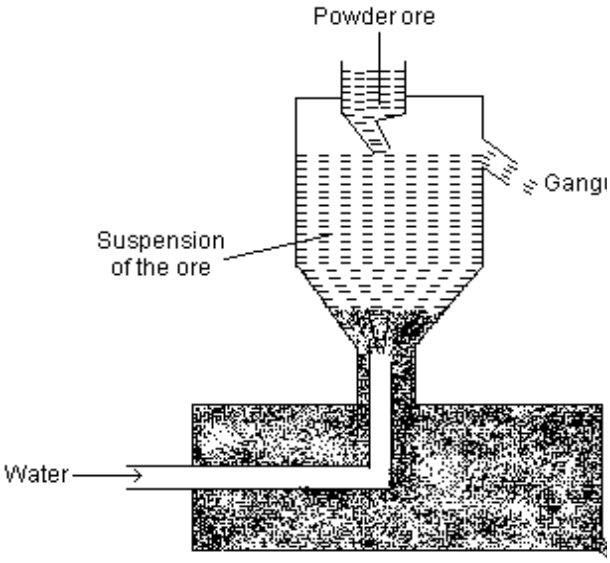
Grinded

Process

Ores are broken into smaller pieces with the help of ***crushers***. There pieces are than converted into find powder with the help of stamp mill. This process is called ***pulverization*** of ore.

Following methods can be applied to remove impurities.

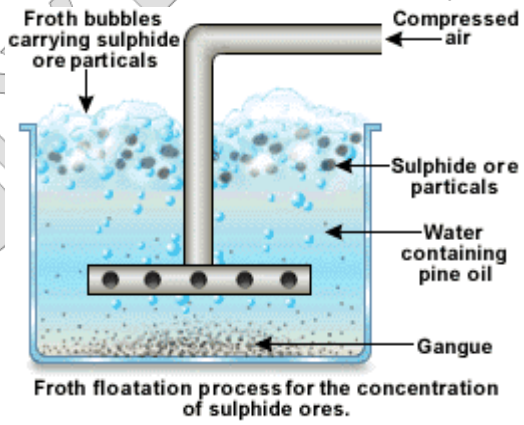
1. ***Hand picking:*** The impurities can be removed mechanically by hand pricking if ore and impurities are differing in size and shape.
2. ***Gravity separation or levigation:***
   This process is done in those cases in which the gangue impurities are then ore. In this process impurities are removed by washing with a stream of water. The heavier ore particles settle low. It is also done in ***wifely tables***.
   This method is applied in case of oxide ore of iron and tine.



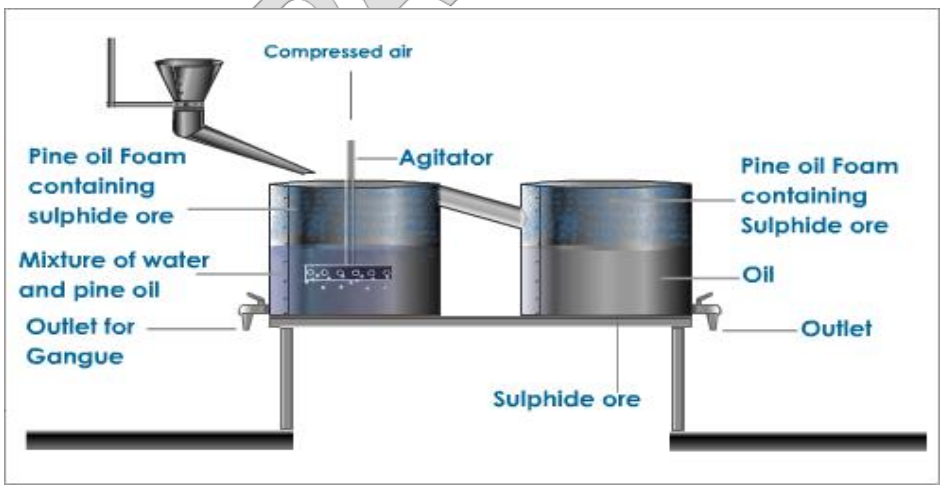
3. ***Froth floatation process:*** It is based on the difference in the welting qualities of the gangue and sulphides are particles by water and oil, where as the ore particles are wetted by water.

Process:

The crushed ore is taken in big tank. It is mixed with H_2O and pine oil (***eucalyptus oil***). Air is blown under pressure into the mixture the oil and air form foam which rise to the surface along with the ***ore*** particles the gangue impurities being wetted by H_2O settle to the bottom of the tank.
The foam is transferred to another tank. It is then repeatedly washed with water to free the particles from impurities.



Froth floatation process for the concentration of sulphide ores.

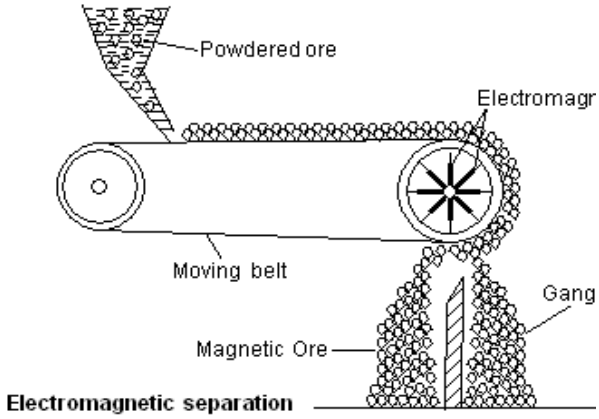


4. ***Magnetic separation :*** It is used in those cases where either the ore or the impurities are magnetic nature.

Example: $Fe(CrO_2)_2$ cremite.

Process:

In this method, the powdered impure ***ore*** is dropped on a leather belt which moves over two rollers one of which has a magnet attached to it. As the ***Ore*** particles roll over the belt. The magnetic component in the ore gets attracted towards the magnet. It gets collected in a heap while the non magnetic component forms a separate heap.



Electromagnetic separation

Extracting metals towards the top of the activity series

Metals that are present high up in the activity series (Na, Ca, Mg etc) are of reactive nature. They cannot be obtained from their compounds by calcinations ***roasting*** or ***reduction***. Because their oxides cannot be reduced with the help of coke because, they are themselves very strong reducing agent.

*These metals are obtained by **electrolytic reduction** method.*

Example: ***A. Electrolytic reduction of NaCl to extract*** Na

Na is isolated from molten NaCl by carrying out the electrolysis. Na is deposited at cathode while chlorine gas is evolved at anode.

$$NaCl\ [molten] \longrightarrow Na^+ + Cl$$

At cathode: $Na^+ + e^- \longrightarrow Na$ *[deposited]*

At Anode: $2Cl^- \longrightarrow Cl_2 + 2e^-$ evolved.

B. Al ***is isolated by the electrolysis of molten*** Al_2O_3 ***(Alumina).***

$$Al_2O_3\ (molten) \longrightarrow 2Al^{3+} + 3O^{2-}$$

At cathode: $Al^{3+} + 3e^- \longrightarrow Al$ *(deposited)*

At anode: $O^{2-} \longrightarrow O + 2e^-$ *(evolved)*

Extraction of metals of moderate reactivity

Metals in the reactivity series are usually present as sulphide or carbonate in nature. It is easier to obtain a metal from its oxides as compared to its sulphides and carbonates. Therefore the metals sulphides and carbonates must be converted into metal oxides. This is done by one of the following processes:

(i) Calcinations (ii) roasting

I. Roasting The sulphide ores are converted into oxide by heating strongly below its melting point in the presence of excess of air.

In this process following changes occurs

(i) Any organic matter if present gets destroyed.

(ii) Impurities of non-metals (sulphur, arsenic, phosphorus) are converted into their volatile oxides.

$S_8 + 8O_2 \longrightarrow 8SO_2$ *(Sulphure –di-oxide)*

$P_4 + 5O_2 \longrightarrow P_4O_{10}$ *(Phosphorous (v) oxide)*

$4As + 3O_2 \longrightarrow 2AS_2O_3$ *(Arsenic-oxide)*

(iii) Sulphides of metals are converted into their oxides.

$2ZnS + 3O_2\ (g) \longrightarrow Zn\ O + 2SO_2\ (g)$

$2PbS + 3O_2\ (g) \longrightarrow 2PbO + 2SO_2\ (g)$

(iv) Small amount of sulphide may be converted into sulphates as a result of roasting.

$ZnS + 2O_2\ (g) \longrightarrow ZnSO_4$

$PbS + 2O_2\ (g) \longrightarrow PbSO_4$

II. Calcination

The process of heating the ore below its melting point in the absence of air is known as ***calcination***. It is applied when ores are metals carbonates.

In this process following changes takes place.

1. Moisture is driven oil and ore become dry.
2. Some hydrated ores decomposed to from anhydrous by losing molecules of water of crystallization.

$Al_2O_3 . 2H_2O \longrightarrow Al_2O_3 + 2H_2O$

(Bauxite)

$Fe_2O_3 . 3H_2O \longrightarrow Fe_2O_3 + 3H_2O$

(Limonite) *(ferric-Oxide)*

3. Carbonates ore decomposed into oxides of metal.

$ZnCO_3 \longrightarrow ZnO + CO_2$

$PbCO_3 \longrightarrow PbO + CO_2$

$CaCO_3 . MgCO_3 \longrightarrow CaO + MgO + 2CO_2$

Conversion of the oxide into metals (Reduction)

The oxides of metals formed in roasting or calculation are reduced to form metal in a number of ways.

(i) Reduction by carbon (smelting)

In this process roasted are [oxide of metal] is mixed with carbon [charcoal, coal or coke] and heated to a temperature above its melting point in ***furnace***.

Smelting is normally done for the extraction of metals like zinc, copper, Fe, Pb, and Sn.

$ZnO + C \longrightarrow Zn + CO$

$SnO_2 + 2C \longrightarrow Sn + 2CO$

$PbO + C \longrightarrow Pb + CO$

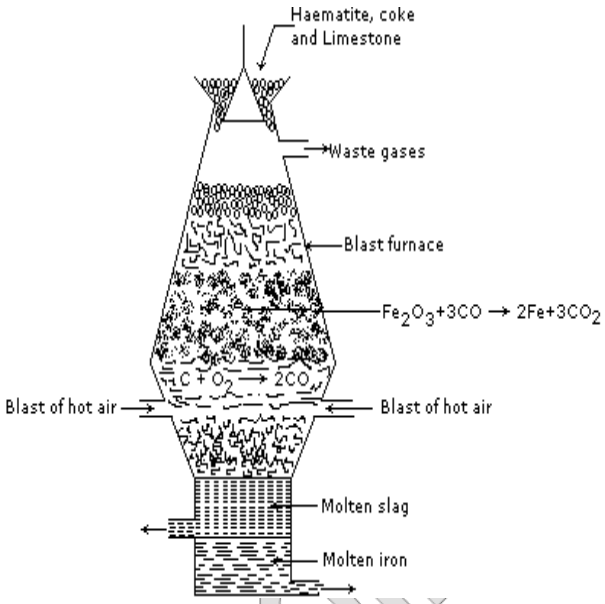
In case of Fe_2O_3

$2C + O_2 \longrightarrow 2CO$

$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$

$2CO + O_2 \longrightarrow 2CO_2$

CO reacts with O_2 and form CO_2 (g)



(ii) Reduction by highly electropositive metals.

It is basically ***metal displacement reaction***. In this process highly electro-positive metals such as Al, Mg can be used to reduce metal oxide like Cr_2O_3 , MnO_2, Fe_2O_3

$$3MnO_2 + 4\,Al \longrightarrow 3Mn + 2Al_2O_3 + \Delta$$

$$Fe_2O_3 + 2Al \longrightarrow 2Fe + Al_2O_3 + \Delta$$

These displacement reactions are highly reactive. The amount of heat evolved is so large that the metals are produces into molten state.

The reaction of **Fe_2O_3** with **Al** used to join railway tracks or cracks of machine part this reaction is known as the ***hermit reaction.***

Extracting metals low in the reactivity series.

Silver (Ag), gold (Au) and platinum (Pt) generally occur in the free or native state. This means that they can be isolated rather easily. Metals like copper (Cu) and mercury (Hg) are comparatively more reactive and occur in combined states. For example, the ore of mercury is cinnabar (HgS) while that of copper is copper glance (Cu_2S). Both are converted into metallic form upon heating in air or oxygen.

$$2HgS\,(s) + 3O_2\,(g) \xrightarrow{Heat} 2HgO\,(s) + 2SO_2\,(g)$$

$$2HgO\,(s) \xrightarrow{Heat} 2Hg\,(l) + O_2\,(g)$$

In extraction of copper,

$$2Cu_2S\,(s) + 3O_2\,(g) \xrightarrow{Heat} 2Cu_2O\,(s) + 2SO_2\,(g)$$

$$2Cu_2O\,(s) + Cu_2S\,(s) \xrightarrow{Heat} 6Cu\,(s) + SO_2\,(g)$$

Refining of metals

The metals produced by various reduction processes are not pure. They contain impurities. There are known as ***crude metals***. Impurities must be removed to obtain pure metal.

The process of getting pure metal from crude metals by a number of methods is known as ***refining of metals***.

The most widely used method for refining impure metal is electrolytic refining.

Electrolytic refining

This method is commonly used for the purification of the metals like Cu, Ag, Zn and Ni.

It is done in an ***electrolytic cell*** in which anode and cathode are used. ***Cathode*** is made up of pure metal. ***Anode*** is made of impure metal. The cell contains solution of ***metal salt*** which acts as an ***electrolyte***.

On passing the electric current through the electrolyte, the pure metal from the anode dissolve into electrolyte.

The equivalent amount of pure metal from the electrolyte is deposited on the cathode. The soluble impurities go into the solution where as the insoluble impurities settle down at the bottom of the anode and are known as ***anode mud***.

Example:

In the electrolytic refining of crud copper impure metal is made ***anode*** and a plate of a pure copper act as the ***cathode***. The electrolyte is the aquatic solution of $CuSO_4$ (copper sulphate). On passing electric current the following changes occur.

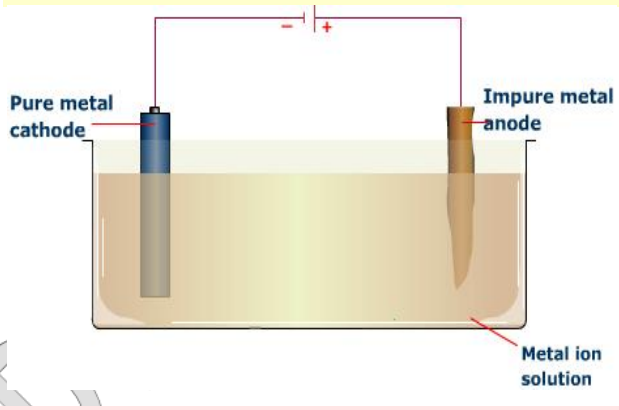
In solution of $CuSO_4$:

$Cu^{2+}(aq)$ $SO_4^{-2}(aq)$ are present

At cathode: $Cu^{2+}(aq) + 2e^- \longrightarrow Cu$ *(deposited)*

At anode: $Cu(s) \longrightarrow Cu^{2}(aq) + 2e^-$

In electro refining the metal from anode does get directly deposited on the cathode.



Corrosion

The process of slow eating up of a metal by the gases and water vapors present in the air due to formation of certain chemical compounds is known as corrosion.

Example:

a. Silver articles become black after some time when exposed to air. This is because it reacts with H_2S (hydrogen sulphide gas) present in the air to form a coating of ***silver sulphide***.

$$Ag + H_2S \longrightarrow \underset{\text{Silver sulphide}}{Ag_2S} + H_2$$

b. Cu reacts with most CO_2 present in the air and slowly loses its shiny brown surface and gain a ***green coat***. This green substance is ***copper carbonate***.

$$\underset{\text{(Most air)}}{2Cu + H_2O + CO_2 + O_2} \longrightarrow \underset{\text{(basic copper carbonate)}}{Cu(OH)_2 .CuCO_3}$$

c. Iron when exposed to most air for a long time acquires a coaling of a brown flaky substance called ***rust.*** This corrosion is called ***rusting***.

$$4Fe + 3O_2 + 2H_2O \longrightarrow 2Fe_2O_3 .2H_2O$$

$$2Fe + \frac{3}{2} O_2 + xH_2O \longrightarrow Fe_2O_3 . xH_2O$$

Factors responsible for corrosion:

1. ***Position of metal in the reactivity series.*** Active metals placed above hydrogen in the reactivity series are easily corroded as compared to the metals which are placed below hydrogen.
2. ***Air and moisture***. The presence of water vapours and gases likeCO_2, SO_2 etc. in air helps the process of corrosion.
3. ***Uneven metal surface***. If the surface of metal is uneven, it will have certain depressions. Water drops will stick in these and take part in the chemical process leading to corrosion.

4. ***Presence of salts***. Presence of salts or electrolytes in water promotes corrosion.
   Example: Rusting of iron is faster in sea water (also called saline water) than in ordinary water or distilled water.

Advantage of corrosion

The metal placed high in the activity series is quite reactive. It combines with the O_2 present in the air to form aluminum oxide Al_2O_3. The oxide-layer gets slowly deposited on the surface of the metal and make it passive to the attack by water, air, acids, alkalies. As a result, the articles made from aluminum are not corroded.

Use: Aluminum foils are used for packing food preparation on cigarettes acts.

Prevention of Corrosion

A number of methods are used to prevent corrosion.

a) ***Painting:*** Paint forms a protective coating on the surface of the articles. As a result, it does not come in contact with the moisture present in air. Even rust proof paints are available these days.
b) ***Oiling and greasing.*** Both protect the surface of metal against moisture and chemicals etc. in addition, the oil and grease prevent the surface from getting scratched.
c) ***Electroplating***. It is very common and effective method to check corrosion or rusting. The surface of iron metal is coating with chromium, nickel or aluminum etc. by electrolysis also called ***electroplating***. They are quite resistant to the attack by both air and water and check corrosion.
   a. ***Galvanization***: If the surface of metal is electroplated by zinc, it is known as galvanization.
   b. ***Tinning.*** If the surface of metal is electroplated by tin, then the process is called tinning.

Galvanization and tinning are very common for checking rusting of iron.

Galvanization is better than tinning: in the reactivity series, the order in which some metals appear is: Zinc (Zn), Iron (Fe) and Tin(Sn). Now, if some scratches appear on the surface of iron article, coating of zinc will be still effective because zinc will be corroded in preference to iron since it is placed above it. However, if same happens in case of tinning then iron will be corroded in preference to tin because, it is more reactive than tin. This means that corrosion will again start. Thus, *galvanization is always better than tinning*.

Alloys

The homogeneous mixture of more than one metal or even metal and non metal is known as ***alloy.***

Procedure:

It is prepared by melting the main constituting metal and then adding definite proportion of the other elements (metal or nonmetal) also in the molten state. The molten mass is stirred uniformly and cooled to form a solid mass. It is an alloy.

Amalgams:

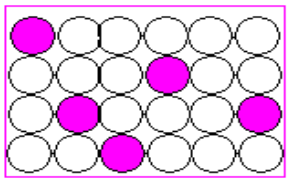
The alloy in which mercury is one of the constituents but not the major one is known as amalgam.

Most of the metals form amalgams with mercury. ***Iron*** and ***platinum*** are the exceptions. Amalgams may be solid or liquid. They are good reducing agents.

Example: sodium amalgam (Na/Hg), magnesium amalgam (Mg/Hg) etc.

Use: Dental alloy used for filling cavities in the teeth contains mercury, silver, tin, copper and a small amount of zinc.

*Alloys are also known as **substitution-solid solution**. Metals have crystal structures which are known as crystal lattices. In the formation of an alloy, certain atoms of a particulars metal present in the lattice are replaced or substituted by the atoms of other metal. As a result, a substitution-solution formed.*



Purpose of alloy making:

Alloy are formed to improve the quality of constituent or metal

- ***Increase hardness***. Alloys are generally harder than the pure metals.
  Example: stain steel and chrome steel are harder than pure iron.
- ***Increase tensile strength***: Alloy steels in general, have greater tensile strength as compared to iron. Chrome steel used in making axels and ball bearings because it has greater tensile strength than iron.
- ***Lower melting point***: Normally alloys have lower melting pints than pure metals.
  Example: solder *(an alloy of tin and lead)* has lower melting point than both the metals.
- ***Modify chemical reactivity***. Alloys are formed to modify the chemical reactivity of metals.
  Example: sodium metal immediately catches fire when added to water. However, sodium amalgam *(an alloy of sodium with mercury)* does not catch fire.
- ***Increase resistance towards corrosion***. Iron easily gets rusted. However different alloy steels normally do not get rusted.

Alloys of Aluminum

Alloy	Percentage Composition	Uses
Magnallium (light and hard)	Al=95%, Mg = 5%,	Pressure cookers, balance beams, some light instruments.
Duralumin (light, strong)	Al = 95%, Cu = 4%, Mg = 0.5%, Mn = 0.5%	Making parts of aeroplanes and automobiles, pressure cookers etc.

Alloys of Copper

Alloy	Percentage Composition	Uses
Bronze (strong , not corrode)	Cu= 90, Sn = 10	For making statues, coins, utensils etc
Brass More malleable , strong)	Cu=80, Zn = 20	For making utensils, parts of machinery, condenser rubes, wires etc.
Gun metal	Cu=90, Sn = 10	For making gun barrels
Bell metal	Cu=80, Sn = 20	For making bells and gongs
German silver	Cu=60, Zn = 20, Ni = 20	For making silver wares, resistance wires.

Phosphor bronze	Cu=95, Sn= 4.8, P = 0.2	For making springs, electric witches.
Monel metal	Cu=30, Ni=67 Fe and Mn=3	For making corrosion resistant pumps and containers for storing acids.

Alloys of Silver

Alloy	Percentage composition	Uses
Coinage silver	Ag=90, Cu=10	For making silver coins
Silver solder	Ag- 63, Cu=30, Zn = 7	For soldering
Dental alloy	Ag=33, Hg= 52, Sn = 12.5, Cu= 2, Zn = 0.5	For filling teeth.

Alloys of tin and lead

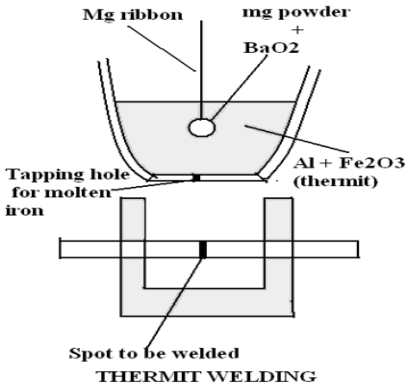
Alloy	Percentage Composition	Uses
Solder (less melting pint)	Pb= 50, Sn = 50	for soldering broken pieces
Type metal	Pb=70, Sb = 20, Sn = 10	For making printing type

Alloys of Iron or Alloy steels (strong and hard)

Alloy	Percentage Composition	Uses
Stainless steel	Fe=73, Cr = 18 Ni = 8 and C = 1	Utensil, cycle and automobile parts Shaving blades, watch cases
Nickel steel	Fe= 98 - 96, Ni = 2 – 4	Cables, automobile and aeroplane parts, armour plates gears and drilling machines
Alnico	Fe=60, Ni = 20, Al= 12, Co = 8	permanent magnets
Chrome steel	Fe = 98, Cr=2	Axels, ball bearings, files and cutting tools.

Thermit welding: Reduction of ferric oxide by aluminum is used in thermit welding. A mixture of ferric oxide and aluminum powder is called thermit (or thermite). It is placed in a crucible made up of silica (SiO_2). It is ignited with the help of magnesium ribbon attached to magnesium powder and barium peroxide which constitute starting plug. Upon ignition, magnesium burns in the presence of barium peroxide (BaO_2) which helps ignition. As a result of heat produced, aluminum present in thermit reduces ferric oxide to iron.

The heat evolved keeps iron in molten state. It escapes from the tapping hole and fills the gap of the broken articles.



THERMIT WELDING

Question with solutions

1. **Compare the properties of a typical metal and a non-metal on the basis of the following. Fill in Column A, B.**

Properties	A Metal	B Non- metal
Electronic configuration	?	?
Nature of oxides	?	?
Oxidizing or reducing action	?	?
Conduction of heat and electricity	?	?

Answer:

Properties	A Metal	B Non- metal
Electronic configuration	Have 1, 2, 3 valence electrons	Have 4, 5, 6, 7 valence electrons
Nature of oxides	Form metallic oxides i.e., basic or amphoreic oxides	Formic acidic oxides or neutral oxides
Oxidizing or reducing action	Metals are donors of electrons $M - e \rightarrow M^+$ Therefore they act as reducing agents	Non-metals generally accept electrons $X + e \rightarrow X^-$ and act as oxidizing agents
Conduction of heat and electricity	Good conductors of heat and electricity	Non-conductors of heat and electricity

2. ***Name the following:***
   1) A molten metal that catches fire in chlorine gas and gives off white fumes
   2) A metal that forms two types of oxides and rusts in moisture; write their formulae also
   3) A metal used in hot water systems
   4) A metal used in long distance cables wires
   5) A metal added to gold to harden it

Answer:

1) The molten metal that catches fire in chlorine gas and gives off white fumes is sodium.
2) The metal that forms two types of oxides and rusts in moisture is iron. The formulas of its oxides are: FeO; Fe_2O_3.
3) The metal used in hot water systems is copper, because it is a good conductor of heat and electricity.
4) The metal used in long distance cables wires is aluminium, because it is a light metal and a very good conductor of electricity.
5) The metal added to gold to harden it is copper.

3. ***A copper plate was dipped in $AgNO_3$ solution. After certain time silver from the solution was deposited on the copper plate. State the reason why it happened. Give the chemical equation of the reaction involved.***

Answer: Copper is more reactive than silver, therefore, it displaces silver from silver nitrate.

$$\underset{\text{Copper}}{Cu} + \underset{\text{Silver nitrate}}{2AgNO_3(aq)} \longrightarrow \underset{\text{Blue}}{Cu(NO_3)_2} + \underset{\text{Silver}}{2Ag}$$

4. ***An element X on reacting with oxygen forms an oxide X_2O. This oxide dissolves in water and turns blue litmus red. State whether element X is metal or a non-metal.***

Answer: Since the oxide turns blue litmus red, therefore X is a non-metal because it is an acidic oxide.

5. ***Metals replace hydrogen from acid, whereas non-metals do not. Why?***

Answer: Non-metals cannot supply electrons to convert H^+ to $H_2(g)$ whereas metals can give electrons to convert.

$$2H^+ + 2e^- \longrightarrow H_2(g).$$

6. ***a) What is the behaviour of magnesium when it is heated and steam is passed over it? Represent the equation. b) Can carbon dioxide react with magnesium?***

Answer: a) When steam is passed over heated magnesium a reaction takes place where 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.

$$Mg_{(s)} + H_2O_{(g)} \Rightarrow MgO_{(s)} + H_{2(g)}$$

b) Ironically, magnesium will even burn in carbon dioxide forming black specks of carbon!

$$2Mg_{(s)} + CO_{2(g)} \Rightarrow 2MgO_{(s)} + C_{(s)}$$

7. ***A zinc rod was kept in a glass container having CuSO4 solution. On examining, it was found that the blue colour of the solution had faded. After few days when the zinc rod was taken out of the solution, a number of small holes were noticed in it. State the reason and give equation of chemical reaction involved.***

Answer:

$$\underset{\text{Blue}}{Zn + CuSO_4} \longrightarrow \underset{\text{colourless}}{ZnSO_4} + Cu$$

Zinc has displaced Cu from $CuSO_4$ solution; therefore holes were noticed on zinc plate.

8. ***Carbonates and sulphide ores are usually converted into oxide ores, why?***

Answer: It is easier to reduce oxide as compared to carbonate and sulphide.

9. ***How does roasting differ from calcination? Give equations.***

Answer:

S.No	Roasting	Calcinations
1.	Roasting is heating ore in the presence of air to oxidize it	Calcinations is heating of the ore, in the absence of air to decompose it or drive volatile matter
2.	Roasting occurs at higher temperatures, higher than the melting point of the ore	Calcinations occurs at temperatures lower than the melting point of the ore

$$\text{Roasting: } 2ZnS + 3O_2 \xrightarrow{\Delta} 2ZnO + 2SO_2$$

$$\text{Calcination: } ZnCO_3 \xrightarrow{\Delta} ZnO + CO_2\uparrow$$

In both cases, the ore becomes porous for easy reduction.Sl. NoRoastingCalcinations1.Roasting is heating ore in the presence of air to oxidize itCalcinations is heating of the ore, in the absence of air to decompose it or drive volatile matter2.Roasting occurs at higher temperatures, higher than the melting point of the oreCalcinations occurs at temperatures lower than the melting point of the ore

10. Why are aluminium containers used to transport nitric acid?

Answer: Aluminium containers are used to transport nitric acid because concentrated HNO_3 renders aluminium passive by forming a thin strong protective oxide layer. This protective layer prevents further reaction with the acid.

11. 1) Why is aluminum used to reduce metal oxides like Fe2O3?
2) What is "Thermite Welding"?

Answer: a) Aluminium is a powerful reducing agent i.e., has great affinity for oxygen. Hence it reduces metallic oxides below it like Fe_2O_3 to metal, with evolution of lot of heat.

2) The property of aluminium to act as a powerful reducing agent and evolving a lot of heat is used in welding broken iron pieces.

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

12. Why is Al obtained only by electrolytic reduction of Alumina (pure)?

Answer: Aluminium metal has a strong affinity for oxygen and the oxide of Al is Al_2O_3, which is a very stable oxide. Hence ordinary chemically reducing agents are not sufficient to reduce Al_2O_3 to Al. Electrolytic reduction is a more powerful reduction method.

13. Name two metals that occur in a free state nature. What is the method of refining them?

Answer: Gold and Platinum are two metals that occur in a free in state in nature. They only need to be cleaned by physical methods.

14. ***Question: Explain the formation of an ionic compound between a metal and a non metal by*** transfer of electrons with Mg as the metal and chlorine as the non-metal to illustrate your answer. Give the reaction that occurs.

Answer: Magnesium, whose atomic number is 12, has 2, 8, 2 configuration. It has two electrons in its valence shell. Chlorine has an atomic number of 17 and an electronic configuration of is 2, 8, 7. It has seven valence electrons. Chlorine is one electron short of the argon configuration (2,8,8) while magnesium has two electrons in excess of the neon configuration (2, 8,). Accordingly, one atom of magnesium will find two atoms of chlorine to transfer its two electrons to (one to each) as shown below:

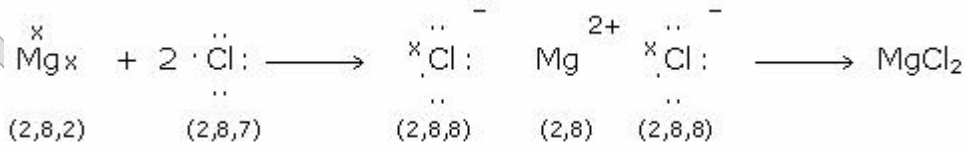
$$\underset{(2,8,2)}{Mg} \longrightarrow \underset{(2,8)}{Mg^{2+}} + 2e^-$$

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

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,



NCERT Solutions

1. Give an example of a metal which (i) is a liquid at room temperature. (ii) can be easily cut with a knife. (iii) is the best conductor of heat. (iv) is a poor conductor of heat.

Ans.: (i) Metal that exists in liquid state at room temperature → Mercury
(ii) Metal that can be easily cut with a knife → Sodium
(iii) Metal that is the best conductor of heat → Silver

(iv) Metals that are poor conductors of heat → Mercury and lead

2. Explain the meanings of malleable and ductile.

Ans.: **Malleable**: Substances that can be beaten into thin sheets are called malleable. For example, most of the metals are malleable. **Ductile**: Substances that can be drawn into thin wires are called ductile. For example, most of the metals are ductile.

3. Why is sodium kept immersed in kerosene oil?

Ans.: Sodium and potassium are very reactive metals and and combine explosively with air as well as water. Hence, they catch fire if kept in open. Therefore, to prevent accidental fires and accidents, sodium is *stored* immersed in kerosene oil.

4. Write equations for the reactions of (i) iron with steam (ii) calcium and potassium with water

(i) $3Fe_{(s)} + 4H_2O_{(g)} \longrightarrow Fe_3O_{4(aq)} + 4H_{2(g)}$

Iron Steam Iron (II,III) oxide Hydrogen

(ii) $Ca_{(s)} + 2H_2O_{(l)} \longrightarrow Ca(OH)_{2(aq)} + H_{2(g)} + Heat$

$2K_{(s)} + 2H_2O_{(l)} \longrightarrow 2KOH_{(aq)} + H_{2(g)} + Heat$

Calcium/ Potassium Water Calcium Hydroxide/ Potassium hydroxide Hydrogen

5. ***Samples of four metals A, B, C and D were taken and added to the following solution one by one. The results obtained have been tabulated as follows.***

Metal	Iron (II) sulphate	Cooper (II) sulphate	Zinc sulphate	Silver nitrate
A.	No reaction	Displacement		
B.	Displacement		No reaction	
C.	No reaction	No reaction	No reaction	Displacement
D.	No reaction	No reaction	No reaction	No reaction

Use the Table above to answer the following questions about metals A, B, C and D.

(i) Which is the most reactive metal? (ii) What would you observe if B is added to a solution of copper (II) sulphate? (iii) Arrange the metals A, B, C and D in the order of decreasing reactivity.

Explanation

A + $FeSO_4$ → No reaction, i.e., A is less reactive than iron

A + $CuSO_4$ → Displacement, i.e., A is more reactive than copper

B + $FeSO_4$ → Displacement, i.e., B is more reactive than iron

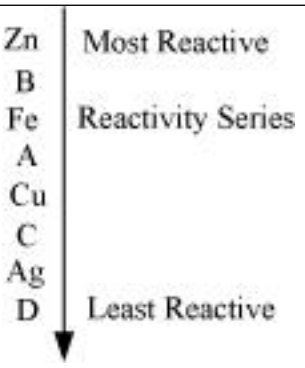
B + $ZnSO_4$ → No reaction, i.e., B is less reactive than zinc

C + $FeSO_4$ → No reaction, i.e., C is less reactive than iron

C + $CuSO_4$ → No reaction, i.e., C is less reactive than copper

C + $ZnSO_4$ → No reaction, i.e., C is less reactive than zinc

C + $AgNO_3$ → Displacement, i.e., C is more reactive than silver



D + $FeSO_4$/$CuSO_4$/$ZnSO_4$/$AgNO_3$ → No reaction, i.e., D is less reactive than iron, copper, zinc, and silver

From the above equations, we obtain:

(i) B is the most reactive metal.

(ii) If B is added to a solution of copper (II) sulphate, then it would displace copper. B + $CuSO_4$ → Displacement

(iii) The arrangement of the metals in the order of decreasing reactivity is:
B > A > C > D

6. Which gas is produced when dilute hydrochloric acid is added to a reactive metal? Write the chemical reaction when iron reacts with dilute H_2SO_4.

Ans.: Hydrogen gas is evolved when dilute hydrochloric acid is added to a reactive metal. When iron reacts with dilute H_2SO_4, iron (II) sulphate with the evolution of hydrogen gas is formed.

$$Fe_{(s)} + H_2SO_{4(aq)} \longrightarrow FeSO_{4(aq)} + H_{2(g)}$$

7. What would you observe when zinc is added to a solution of iron (II) sulphate? Write the chemical reaction that takes place.

Ans.: Zinc is more reactive than iron. Therefore, if zinc is added to a solution of iron (II) sulphate, then it would displace iron from the solution.

$$Zn_{(s)} + FeSO_{4(aq)} \longrightarrow ZnSO_{4(aq)} + Fe_{(s)}$$

8. (i) Write the electron-dot structures for sodium, oxygen and magnesium. (ii) Show the formation of Na_2O and MgO by the transfer of electrons. (iii) What are the ions present in these compounds?

Ans.: (i) The representation of elements with valence electrons as dots around the elements is referred to as electron-dot structure for elements.

(a) Sodium (2, 8, 1) = $\dot{Na}$

(b) Oxygen (2, 6) = $:\ddot{O}:$

(c) Magnesium (2, 8, 2) = $\dot{M}g$

(ii) $Na + Na + \ddot{O} \longrightarrow (Na^+)_2[:\ddot{O}:]^{2-}$

$Mg + \ddot{O} \longrightarrow (Mg^{2+})[:\ddot{O}:]^{2-}$

(iii) The ions present in Na_2O are Na^+ and O^{2-} ions and in MgO are Mg^{2+} and O^{2-} ions.

9. Why do ionic compounds have high melting points?

Ans.: Ionic compounds have strong electrostatic forces of attraction between the ions. Therefore, it requires a lot of energy to overcome these forces. That is why ionic compounds have high melting points.

10. Name two metals which are found in nature in the free state.

Ans.: The metals at the bottom of the reactivity series are mostly found in free state. For example: gold, silver, and platinum.

11. Metallic oxides of zinc, magnesium and copper were heated with the following metals.

Metal	Zinc	Magnesium	Copper
Zinc oxide	-	-	-
Magnesium oxide	-	-	-
Copper oxide	-	-	-

12. In which cases will you find displacement reactions taking place?

Ans.:

Metal	Zinc	Magnesium	Copper
Zinc oxide	No reaction	Displacement	No reaction

Magnesium oxide	No reaction	No reaction	No reaction
Copper oxide	Displacement	Displacement	No reaction

13. In the electrolytic refining of a metal M, what would you take as the anode, the cathode and the electrolyte?

Ans.: In the electrolytic refining of a metal M:

Anode → Impure metal M, Cathode → Thin strip of pure metal M

Electrolyte → Solution of salt of the metal M

14. Give reasons

a. Platinum, gold and silver are used to make jewellery.

b. Sodium, potassium and lithium are stored under oil.

c. Aluminium is a highly reactive metal, yet it is used to make utensils for cooking.

d. Carbonate and sulphide ores are usually converted into oxides during the process of extraction.

Ans.: (a) Platinum, gold, and silver are used to make jewellery because they are very lustrous. Also, they are very less reactive and do not corrode easily.

(b) Sodium, potassium, and lithium are very reactive metals and react very vigorously with air as well as water. Therefore, they are kept immersed in kerosene oil in order to prevent their contact with air and moisture.

(c) Though aluminium is a highly reactive metal, it is resistant to corrosion. This is because aluminium reacts with oxygen present in air to form a thin layer of aluminium oxide. This oxide layer is very stable and prevents further reaction of aluminium with oxygen. Also, it is light in weight and a good conductor of heat. Hence, it is used to make cooking utensils.

(d) Carbonate and sulphide ores are usually converted into oxides during the process of extraction because metals can be easily extracted from their oxides rather than from their carbonates and sulphides

15. You must have seen tarnished copper vessels being cleaned with lemon or tamarind juice. Explain why these sour substances are effective in cleaning the vessels.

Ans.: Copper reacts with moist carbon dioxide in air to form copper carbonate and as a result, copper vessel loses its shiny brown surface forming a green layer of copper carbonate. The citric acid present in the lemon or tamarind neutralises the basis copper carbonate and dissolves the layer. That is why, tarnished copper vessels are cleaned with lemon or tamarind juice to give the surface of the copper vessel its characteristic lustre.

16. A man went door to door posing as a goldsmith. He promised to bring back the glitter of old and dull gold ornaments. An unsuspecting lady gave a set of gold bangles to him which he dipped in a particular solution. The bangles sparkled like new but their weight was reduced drastically. The lady was upset but after a futile argument the man beat a hasty retreat. Can you play the detective to find out the nature of the solution he had used?

Ans.: He must have dipped the gold metal in the solution of aqua regia – a 3:1 mixture of conc. HCl and conc. HNO_3. Aqua regia is a fuming, highly corrosive liquid. It dissolves gold in it. After dipping the gold ornaments in aqua regia, the outer layer of gold gets dissolved and the inner shiny layer appears. That is why the weight of gold ornament reduced.