Chemistry (class X)
Chapter - 3 (Metals and Non-metals)

Introduction

There are 115 chemical elements known at present. On basis of their properties, all the elements can be divided into two main groups : metal and non-metals.

Malleable mean which can be beaten with a hammer to form thin sheets (without breaking). **Ductile** mean which can be stretched (or drawn) to form thin wire. And **brittle** mean which breaks into pieces on hammering or stretching.

METALS

Metals are the elements that conduct heat and electricity and are malleable and ductile. Metals are also lustrous (shiny), hard, strong, heavy and sonorous (which make ringing sound when struck). Examples: Iron, Aluminium, Silver, Gold, Platinum, Zinc, Tin Mercury, Sodium, Potassium, Calcium and Magnesium. All the metals are solids, except mercury which is a liquid metal.

Physical Properties of metal

(i) Malleability : The property of metals due to which they can be beaten into thin sheets is called malleability. Eg (i) Aluminium foil is used to wrap food items (ii) Silver foil is used to decorate sweets. Gold and Silver are the most malleable metals.

(2) Ductility : The ability of metals to be drawn into thin wires is called ductility. Eg. Copper wires are used as electric wires. Gold is the most ductile metal. A wire of about 2 km length can be drawn from one gram of gold.

(3) Sonorous : The property of a metal due to which it produces a ringing sound on striking. Eg bells are made of metals.

(4) Lustre : The property of a metal due to which it has a shiny appearance. Eg. Gold and silver are used to make jwellery.

(5) Conductivity : Metals are good conductors of heat and electricity and have high melting point.

Non-Metals

Non-metals are the elements that do not conduct heat and electricity, and are neither malleable nor ductile. They are brittle. Non-metals are not lustrous (not shiny), they have dull appearance. Non-metals are generally soft and not strong. They are light substances and non-sonorous. Examples : Carbon, Sulphur, Phosphorus, Silicon, Hydrogen, Oxygen, Nitrogen, Chlorine, Bromine, Iodine, Helium, Neon and Argon.

Some Exception of metals and and non-metal properties

(i) All metals except mercury exists as solid at room temperature.

(ii) Metals have high melting point but Gallium and Caesium have low melting points. These two metals will melt if you keep them on your palm.

(iii) Iodine is a non-metal but it is lustrous

(iv) Carbon is a non-metal that can exist in different forms. Each form is called an allotrope. Diamond, an allotrope of carbon, is hardest natural substance known and has a very high melting and boiling point. Graphite, another allotrope of carbon, is a conductor of electricity.

(v) Alkali metals (Lithium, sodium, potassium) are so soft that they can be cut with a knife. They have low densities and low melting points.

Chemical Properties of Metals.

1) Reaction of Metals with Oxygen (of Air)

Almost all metals combine with oxygen to form metal oxides.

Metal + Oxygen (from air) ⟶ Metal oxide (Basic oxide)

metal oxide, being basic, turns red litmus solution blue.

The vigour (strength) of reaction with oxygen depends on the chemical reactivity of metal. Some metals react with oxygen even at room temperature, some react on heating, whereas still others react only on strong heating. Examples:

(i) Sodium metal reacts with oxygen of air at room temperature to form a basic oxide called sodium oxide: $4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$

Potassium metal (K) also reacts with the oxygen (O_2) of air at room temperature to form a basic oxide, called potassium oxide (K_2O).

Potassium and sodium metal are stored under kerosene oil to prevent their reaction with oxygen, moisture and carbon dioxide of air (so as to protect them). Potassium and sodium are very reactive metals.

Another metal which is very reactive is Lithium (Li) which also stored under kerosene oil.

$2Cu + O_2 \longrightarrow 2CuO$

Copper (II) oxide

$4Al + 3O_2 \longrightarrow 2Al_2O_3$

(Aluminium oxide)

Some metals oxides, such as aluminium oxide, zinc oxide etc, show both acidic as well as basic behaviour. Such metals oxides which react with both acids and well as bases to produce salts and water are known as amphoteric oxides.

Aluminium oxide reacts with following manner with acids and bases :-

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

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

(Sodium aluminate)

Most of the metals oxides are insoluble in water. But some of the metal dissolve in water to for alkalis. Sodium oxide and potassium oxide dissolve in water to produce alkalis as follows :

$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$

$K_2O(s) + H_2O(l) \longrightarrow 2KOH(aq)$

Solution of Sodium oxide and potassium oxide in water turns red litmus to blue.

(ii) Magnesium metal does not react with oxygen at room temperature.

But on heating, magnesium metal burn in air giving intense heat and light & form a basic oxide called magnesium oxide (which is white powder):

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

Since heat is required for the reaction of magnesium with oxygen, it mean magnesium is less reactive than sodium (or potassium).

Magnesium oxide dissolves in water partially & form magnesium hydroxide solution:

$$MgO(s) + H_2O(l) \longrightarrow Mg(OH)_2(aq)$$
(Base)

(Base — turn red litmus solution to blue).

(iii) Aluminium metal: burn in air, on heating, & form aluminium oxide:

$$4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$$
(Amphoteric oxide)

Those metal oxides which show basic as well as acidic behaviour are known as amphoteric oxides.
Aluminium and zinc form amphoteric oxides.
Amphoteric oxides react with both, acids as well as bases & form salts and water.

(iv) **Zinc metal** burn in air only on strong heating to form zinc oxide:

$$2Zn(s) + O_2(g) \longrightarrow 2ZnO(s)$$

(Amphoteric oxide)

(v) **Iron metal**: does not burn in air even on strong heating. But iron filings (small particles of iron) burn vigorously when sprinkled in the flame of a burner.

$$3Fe(s) + 2O_2(g) \longrightarrow Fe_3O_4(s)$$

(Iron) — iron (II, III) oxide.

(vi) **Copper metal**: also does not burn in air even on strong heating. Copper reacts with oxygen of air on prolonged heating to form a black substance copper (II) oxide:

$$2Cu(s) + O_2(g) \longrightarrow 2CuO(s)$$

(Copper (II) oxide)

(vii) **Silver and gold** do not react with oxygen even at higher temperatures.

Reaction of Metals with Water

Metal react with water and produce a metal oxide and hydrogen gas. Metal oxides that are soluble in water dissolve in it to further form metal hydroxide. But all metal do not react with water.

Metal + water ⟶ Metal oxide + Hydrogen

Metal oxide + water ⟶ Metal hydroxide.

(i) Metals like potassium and sodium react violently with cold water. The reaction is so violent and exothermic that the evolved hydrogen immediately catches fire.

$$2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g) + \text{heat energy}$$

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g) + \text{heat energy}$$

(ii) The reaction of calcium with water is less violent. The heat evolved is not sufficient for hydrogen to catch fire.

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

Calcium starts floating because the bubbles of hydrogen gas formed stick to the surface of the metal.

(iii) Magnesium does not react with cold water. It reacts with hot water to form magnesium hydroxide and hydrogen. It also starts floating due to the bubbles of hydrogen gas sticking to its surface.

$$Mg(s) + 2H_2O(l) \xrightarrow{\text{(Hot)}} Mg(OH)_2 + H_2(g)$$

(iv) Metals like aluminium, iron and zinc do not react either with cold or hot water. But they react with steam to form the metal oxide and hydrogen.

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

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

(v) Metals such as lead, copper, silver and gold do not react with water at all.

Summary:

(a) When metal react with cold water, product formed are metal oxide and hydrogen gas.

Metal + Cold water ⟶ Metal oxide + Hydrogen

(b) When metal react with hot water, product formed are metal hydroxide and hydrogen gas

Metal + Hot water ⟶ Metal hydroxide + Hydrogen

(c) When metal reacts with steam, then products formed are metal oxide and hydrogen gas:

Metal + Steam ⟶ Metal oxide + Hydrogen.

Reaction of Metals with Acids.

Metal react with acids to give a salt and hydrogen gas.

Metal + Dilute acid ⟶ Salt + Hydrogen.

(i) $2Na(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2(g)$
(Hydrochloric acid) (Sodium chloride)
This reaction is very reactive.

(ii) $Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$
(Magnesium chloride)

(iii) $2Al(s) + 6HCl(aq) \longrightarrow 2AlCl_3(aq) + 3H_2(g)$
(Aluminium chloride)

(iv) $Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$
(Zinc chloride)

(v) $Fe(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2(g)$
(Iron) (Iron (II) chloride)

(vi) Copper does not react with HCl or H_2SO_4 (Sulphuric acid) at all. This shows that copper is even less reactive that iron
$Cu(s) + HCl(aq) \longrightarrow$ No reaction.

Hydrogen gas is not evolved when a metal reacts with nitric acid (HNO_3). It is because HNO_3 is a strong oxidising agent. It oxidises the H_2 produced to water and itself gets reduced to any of the Nitrogen oxides (N_2O, NO, NO_2)

Aqua - Regia

Aqua - regia is a freshly prepared mixture of 1 part of concentrated Nitric acid and 3 parts of concentrated hydrochloric acid. Thus, the ratio of conc. HNO_3 and conc. HCl in aqua - regia is 1:3.
Aqua - regia is a highly corrosive, fuming liquid.

Aqua - regia can dissolve all metals. It can even dissolve gold and platinum metals.

Reaction of metals with solution of other metals Salts?

Reactive metals can displace less reactive metals from their compounds in solution or molten form.

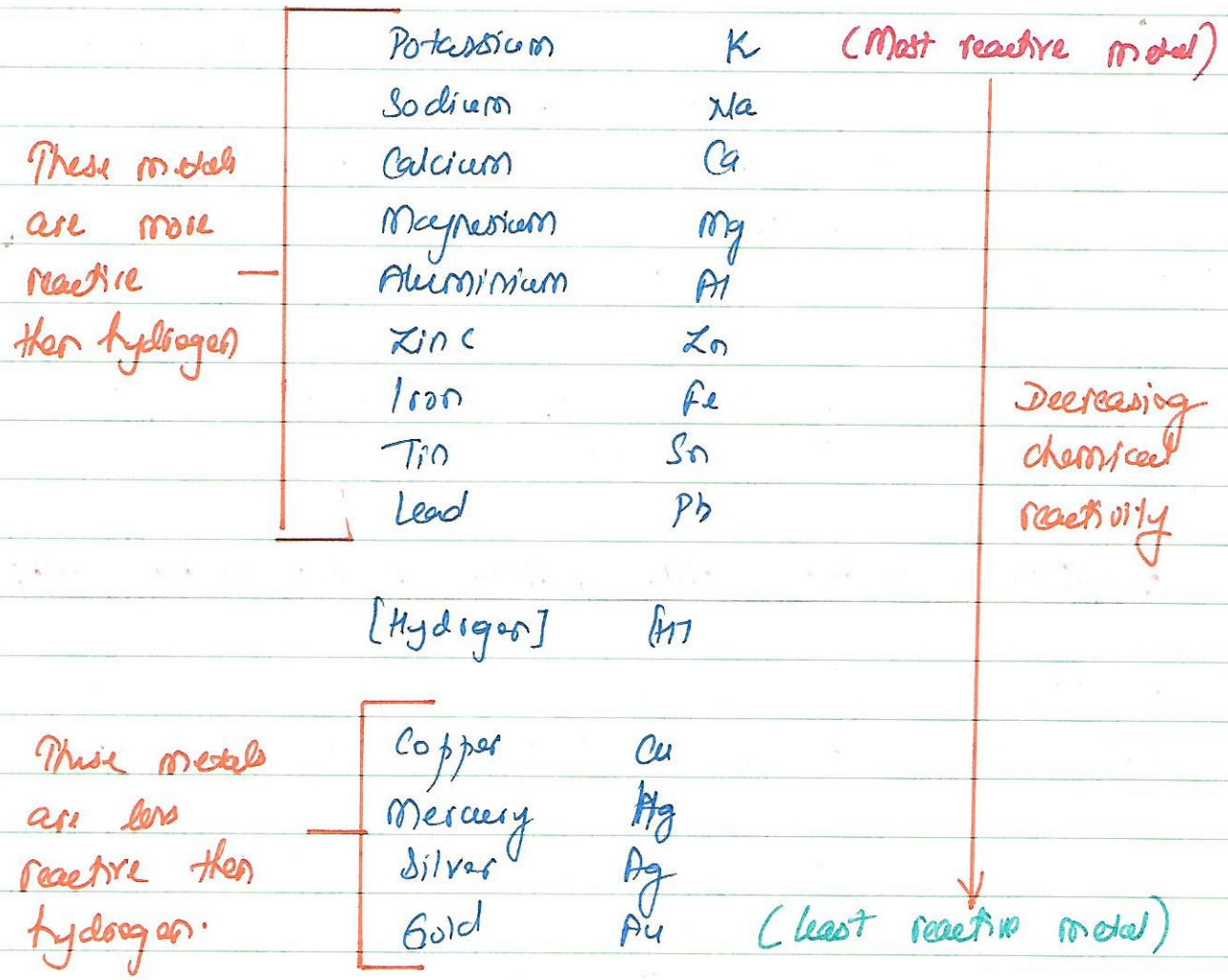
Metal A + Salt solution of B ⟶ Salt solution of A + Metal B
(Displacement reaction.
$CuSO_4(aq) + Zn(s) \longrightarrow ZnSO_4(aq) + Cu(s)$

The Reactivity Series :

The arrangement of metals in a vertical column in order of decreasing reactivities is called reactivity series of metals (or activity series of metals).
In reactivity series, the most reactive metal is placed at the top whereas the least reactive metal is placed at the bottom.

A more reactive metal displaces a less reactive metal from its salt solution.

Reactivity Series (or Activity Series) of metals.



Though hydrogen is not a metal but even then it has been placed in the reactivity series of metals. This is due to the fact that like metals, hydrogen also loses electrons and form positive ion, H^+.

Questions

(1) Give an example of a metal which

(i) is a liquid at room temperature — Mercury

(ii) can be easily cut with a knife — Sodium

(iii) is a best conductor of heat - Silver

(iv) is a poor conductor of heat - Mercury and lead.

(2) Explain the meaning of malleable and ductile.

Ans. Refer * notes.

(3) Why is sodium kept immersed in kerosene oil.?

Ans: Sodium is a highly reactive metal. It can catch fire if kept in open. Hence to protect it and prevent accidental fire, it is stored under kerosene oil.

(4) In spite of being highly reactive, aluminium is still used for making utensils.

Ans: Due to high reactivity of aluminium, a thin layer of aluminium oxide is formed on it. This protective layer of aluminium oxide prevents the metal from further corrosion. Hence, aluminium can be used for making utensils. This process is anodising.

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

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

(Steam) Iron (II,III)

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

$2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g) + Heat.$

(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.

$Fe(s) + H_2SO_4(aq) \longrightarrow FeSO_4(aq) + H_2(g)$

(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) What are amphoteric oxides? Give two examples.

ANS - Refer to notes.

9) Name two metals that start floating after sometime when immersed in water and explain why they do so.

ANS: Calcium (Ca) and Magnesium (Mg). They start floating because the bubbles of hydrogen gas which are formed during the reaction stick to the surface of the metal.

10) Name the metal which reacts with a very dilute HNO_3 to evolve hydrogen gas.

ANS: Manganese (Mn) or Magnesium (Mg)

How do metals and non-metals react?

When metal react with non metals, they form ionic compounds (which contain) ionic bonds). On the other hand, when non-metals react with other non-metals they form covalent compounds (which contain covalent bonds).

Metals, however, do not react with other metals.

Formation of Sodium Chloride (NaCl)

Na — 2, 8, 1 (11) , Cl — 2, 8, 7 (17)

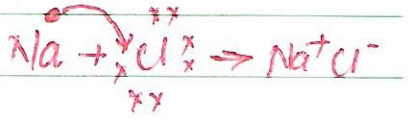
Sodium atom has one electron in its outermost shell. It loses the electron form its outermost shell and becomes a stable octet. The nucleus of this atom still has 11 protons but the number of electrons has become 10, so there is a net positive charge giving us a sodium cation Na^+.

Chlorine has 7 electrons in outermost shell and it requires one more electron to complete its octet.

If sodium and chlorine were to react, the electron lost by sodium could be taken up by chlorine. After gaining an electron, the chlorine atom gets a unit negative charge because its nucleus has 17 protons and there are 18 electrons in its K, L and M shells. This gives us a chlorine anion Cl^-.

$Na \rightarrow Na^+ + e^-$ (2,8,1) (2,8)

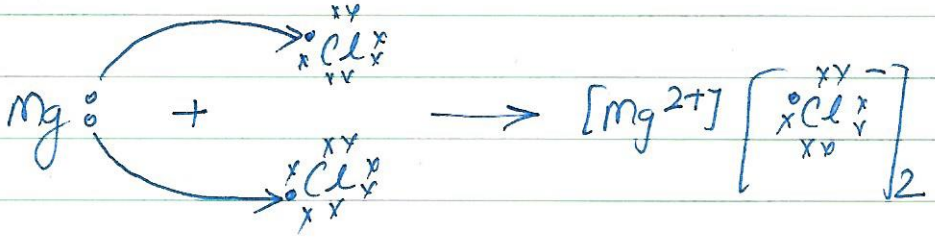
$Cl + e^- \rightarrow Cl^-$ (2,8,7) (2,8,8)



formation of magnesium chloride ($MgCl_2$)

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

$Cl + e^- \longrightarrow Cl^-$ (2,8,7 → 2,8,8)



The compounds formed in this manner by the transfer of electrons from a metal to a non-metal are known as ionic compounds or electrovalent compounds.

IONS

An ion is an electrically charged atom (or group of atoms). Examples: Na^+, Mg^{2+}, Cl^-, O^{2-}.

An ion is formed by the loss or gain of electron by an atom, so it contains an unequal number of electrons and protons.

A positively charged ion is known as cation.

A negatively charged ion is known as anion.

Chemical Bond

When atoms of the elements combine to form molecules, a force of attraction is developed between the atoms (or ion) which hold them together. The force which links the atoms (or ion) in a molecule is called a chemical bond.

Noble Gases: In noble gases the electron arrangement in their atom are very stable which do not allow the outermost electrons to take part in chemical reaction. So called inert gases.

Types of Chemical Bond

There are two types of chemical bonds:
(i) Ionic Bond
(ii) Covalent bond.

IONIC BOND

The chemical bond formed by the transfer of electron from one atom to another is known as an ionic bond.

Ionic bond is formed when one of the atoms can donate electrons to achieve the inert gas electron configuration, and the other needs electrons to achieve the inert gas electron configuration.

When a metal reacts with a non-metal, transfer of electron takes place from metal atoms to the non-metal atoms, and an ionic bond is formed.

Examples: (i) formation of Sodium chloride (NaCl)
(ii) formation of Magnesium chloride ($MgCl_2$)

Ionic Compounds.

The compounds containing ionic bonds are known as ionic compounds. They are formed by transfer of electrons from one atom to another. The ionic compounds are made up of positively charged ion (cation) and negatively charged ion (anion). The ionic compounds consists of ion and not molecules.

Some ionic Compounds :

Sodium chloride	NaCl	Na^+ & Cl^-
Potassium chloride	KCl	K^+ & Cl^-
Magnesium chloride	$MgCl_2$	Mg^{2+} & Cl^-
Sodium oxide	Na_2O	Na^+ & O^{2-}
Calcium oxide	CaO	Ca^{2+} & O^{2-}

Properties of Ionic Compound

(i) Physical Nature : Ionic compounds are solids and are somewhat hard because of the strong force of attraction between the positive and negative ion. These compounds are generally brittle and break into pieces when pressure is applied.

(ii) Melting and Boiling Points : Ionic compounds have high melting and boiling points. This is because a considerable amount of energy is required to break the strong inter-ionic attraction.

(iii) Solubility : Electrolent compounds are generally soluble in water and insoluble in solvents such as kerosene, petrol etc.

(iv) Conduction of Electricity : Ionic compounds in the solid state do not conduct electricity because movement of ions in the solid is not possible due to their rigid structure. But ionic compounds conduct electricity in the molten state. This is possible in the molten state since the electrostatic forces of attraction between the oppositely charged ions are overcome due to heat. The ions freely and conducts.

Covalent Bond

The chemical bond formed by the sharing of electron between two atoms is known as a covalent bond. The sharing of electron takes place in such a way that each atom in the resulting molecule gets the stable electron arrangement of the inert gas. Atoms share only their outermost electron in the formation of covalent bonds.

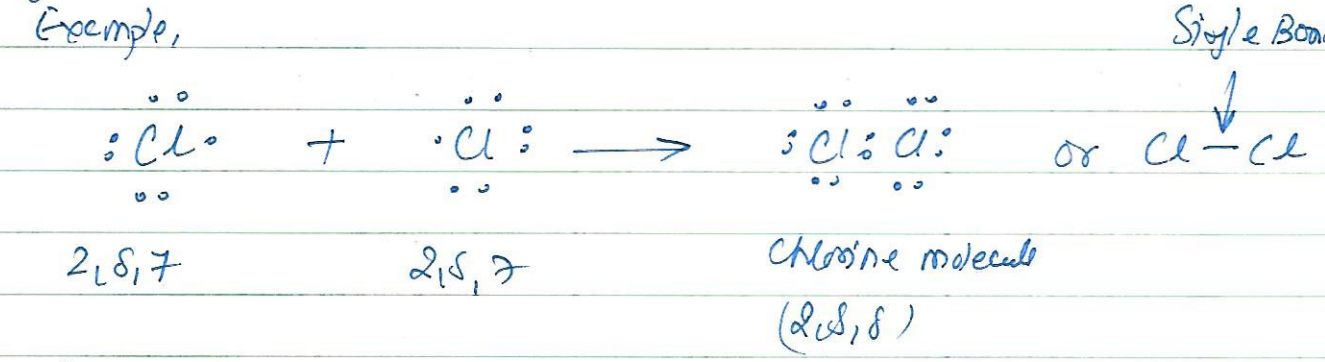
Whenever a non-metal combines with another non-metal, sharing of electron takes place between their atoms and a covalent bond is formed. Example, hydrogen is a non-metal and chlorine is also a non-metal, so when hydrogen combines with chlorine to form hydrogen chloride, HCl, sharing of electron takes place.

Covalent bond can be also formed between two atoms of the same non-metal. Example, two chlorine atoms combine together by the sharing of electron to form a chlorine molecule, Cl_2.

Single Covalent Bond

A single bond is formed by the sharing of one pair of electron between two atoms.

Example,



Formation of hydrogen molecule :

$H\cdot + \cdot H \longrightarrow H:H$ or $H-H$

Formation of HCl : $H\cdot + \cdot\ddot{Cl}: \longrightarrow H:\ddot{Cl}:$ or $H-Cl$

Formation of methane : CH_4

$\cdot\dot{C}\cdot + 4\cdot H \longrightarrow H:\overset{H}{\overset{\cdot\cdot}{C}}:H$ or $H-\overset{H}{\overset{|}{C}}-H$ (with H below)

Double Bond

A double bond is formed by the sharing of two pairs of electrons between two atoms.

Formation of O_2 molecule : O_2

$:\ddot{O}: + :\ddot{O}: \longrightarrow :\ddot{O}::\ddot{O}:$ or $O=O$ (Double Bond)

Two pairs of electrons are shared.

Formation of CO_2 : & C_2H_4

$:C: + 2:\ddot{O}: \longrightarrow :\ddot{O}::C::\ddot{O}:$ or $O=C=O$

$2:C: + 4\cdot H \longrightarrow$ (H, H, C::C, H, H) or $H_2C=CH_2$

Triple Bond

A triple bond is formed by the sharing of three pair of electron between two atoms.

Formation of N_2 molecule: Triple Bond

$$:\dot{N}: + :\dot{N}: \longrightarrow :N:::N: \text{ or } :N \equiv N:$$

Three pair of electrons are shared.

Formation of Ethyne Molecule, C_2H_2 Triple Bond

$$2 \cdot \dot{C}: + 2 \cdot H \longrightarrow H:C:::C:H \text{ or } H-C \equiv C-H$$

Covalent Compounds.

The compounds containing covalent bonds are known as covalent compounds. Covalent compounds are formed by the sharing of electron between atoms.

Examples, CH_4, C_2H_6, C_2H_4, C_2H_2, H_2O, NH_3, HCl, CO_2

Properties of covalent compounds:

(1) Covalent compounds are usually liquids or gases. Only some of them are solids. Example: alcohol, ether, benzene, carbon tetrachloride and bromides. or liquid.
Glucose, cane sugar, urea, naphthalene and iodine are solid covalent compounds.

(2) Covalent compounds have usually low melting points and low boiling points.
Covalent compounds are made up of electrically neutral molecules. So, the force of attraction between the molecules of a covalent compound is very weak. Only a small amount of heat energy is required to break these weak molecular forces, due to which covalent compounds have low melting points and low boiling points.

(3) Covalent compounds are usually insoluble in water but they are soluble in organic solvents.

(4) Covalent compounds do not conduct electricity.
Covalent compounds do not conduct electricity because they do not contain ions.

Ionic Compounds	Covalent Compounds
1) Usually crystalline solid	1) Usually liquid or gases
2) High melting and boiling point	2) low melting and boiling point.
3) Conduct electricity	3) Do not conduct electricity
4) Usually soluble in water	4) Usually insoluble in water
5) Insoluble in organic solvents Like alcohol, ether acetone etc.	5) Soluble in organic solvents.

Questions

1) (i) Write the electron-dot structure for sodium, oxygen and magnesium.

Ans

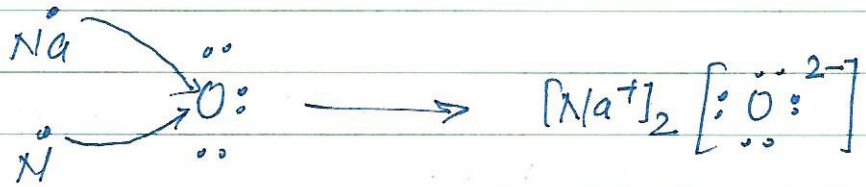
(a) Sodium (2,8,1) – $Na^{\cdot}$

(b) Oxygen (2,6) – $:\ddot{O}:$

(c) Magnesium (2,8,2) – $\ddot{Mg}$

(2) Show the formation of Na_2O and MgO by the transfer of electron.

Ans)



(3) Why do ionic compounds have high melting points?

Ans - Refer to notes.

(4) Which of the following pairs will give displacement reaction?

(a) NaCl solution and copper metal

(b) $MgCl_2$ sol and aluminium metal

(c) $FeSO_4$ soln and silver metal

(d) $AgNO_3$ solution and copper metal.

Ans) (d) $AgNO_3$ solution and copper metal.

(5) Which method is suitable for preventing an iron frying pan from rusting?

Ans) Applying a coating of zinc.

(6) An element reacts with oxygen to give a compound with a high melting point. This compound is also soluble in water. Name the likely element.

Ans) The element is likely to be calcium.

(7) Food cans are coated with tin and not zinc because.

Ans) Zinc is more reactive than tin.

(8) You are given a hammer, a battery, a bulb, wires and a switch.

(a) How could you use them to distinguish between samples of metals and non-metals.?

(b) Assess the usefulness of these tests in distinguishing between metals and non-metals.

Ans) a) With hammer, we can beat the sample and if it can be beaten into thin sheets (that is, it is malleable), then it is a metal otherwise a non-metal. Similarly, we can use the battery, bulb, wires and a switch to set up a circuit with sample. If the sample conducts electricity, then it is a metal otherwise a non-metal.

(b) The above tests are useful in distinguishing between metals and non-metals as these are based on the physical properties. No chemical reactions are involved.

(9) Name two metals which will displace hydrogen from dilute acids, and two metals which will not.

Ans) Metals that are more reactive than hydrogen displaces it from dilute acids. Exam: Na and K. Metals that are less reactive than hydrogen do not displace it. Example: copper and silver.

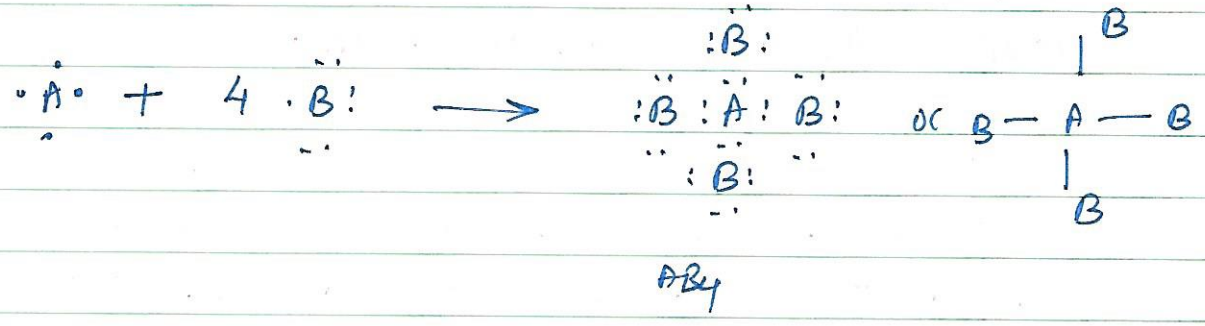
(10) In the formation of compound AB, atoms of A lost one electron each while atoms of B gained one electron each. What is the nature of bond in AB? Predict the two properties of AB.

Ans) Here, atoms of A lose electron whereas the atom of B gains electron. This means that there is a transfer of electron from atoms of A to atom of B. Now, the bond formed by the transfer of electron is called ionic bond. So, the nature of bond in the compound AB is ionic.
Two properties of ionic compound AB will be:
(i) It will be soluble in water
(ii) It will conduct electricity when dissolved in water or melted

(11) An element 'A' has 4 electron in the outermost shell of its atom and combines with another element 'B' having 7 electron in the outermost shell of its atom. The compound formed does not conduct electricity. What is the nature of the chemical bond in the compound? Give the electron-dot structure of its molecule.

Ans) The atom of A has 4 valence electron so it needs 4 more electron to achieve the stable, 8-electron configuration in the outermost shell. The atom B has 7

valence electrons, so it needs 1 more electron to complete the 8-electron structure. Since both the reacting atoms need electrons to achieve the inert gas electron arrangements, they will combine by the sharing of electrons and form covalent bonds. Thus, the nature of chemical bond present in the compound is "covalent bond". The presence of covalent bonds in the compound is confirmed by the fact that the compound does not conduct electricity (only ionic compounds containing ionic bonds conduct electricity). We will now give the electron-dot structure of a molecule of the compound formed.



AB_4

(12) Give the formula of the chlorides of the elements A and B having atomic numbers of 6 and 11 respectively. Will the properties of the two chlorides be similar or different. Explain

Sol) A – 2, 4 (6) B = 2, 8, 1 (11)

Since valency of A is 4 and chlorine is 1, so one atom of element A will share its four electrons with 4 atoms of Cl to form covalent bond (compound) having formula ACl_4.

Atom B has only 1 electron in its outermost shell which it can give to chlorine ion and form an cation B^+. Now, the ions B^+ and Cl^- combine to give an ionic compound having the formula BCl.

Occurance of metals

- The earth's crust is major source of metals.
- Seawater also contains some soluble salts such as sodium chloride, magnesium chloride etc.
- The elements or compounds, which occur naturally in the earth crust are known as minerals.
- At some places, minerals contain a very high percentage of a particular metal and metal can be profitably extracted from it. These minerals are called ores.

Extraction of Metals

To obtain a metal from its ore is called the extraction of metals. The various processes involved in the extraction of metals from their ore, and refining are known as metallurgy.

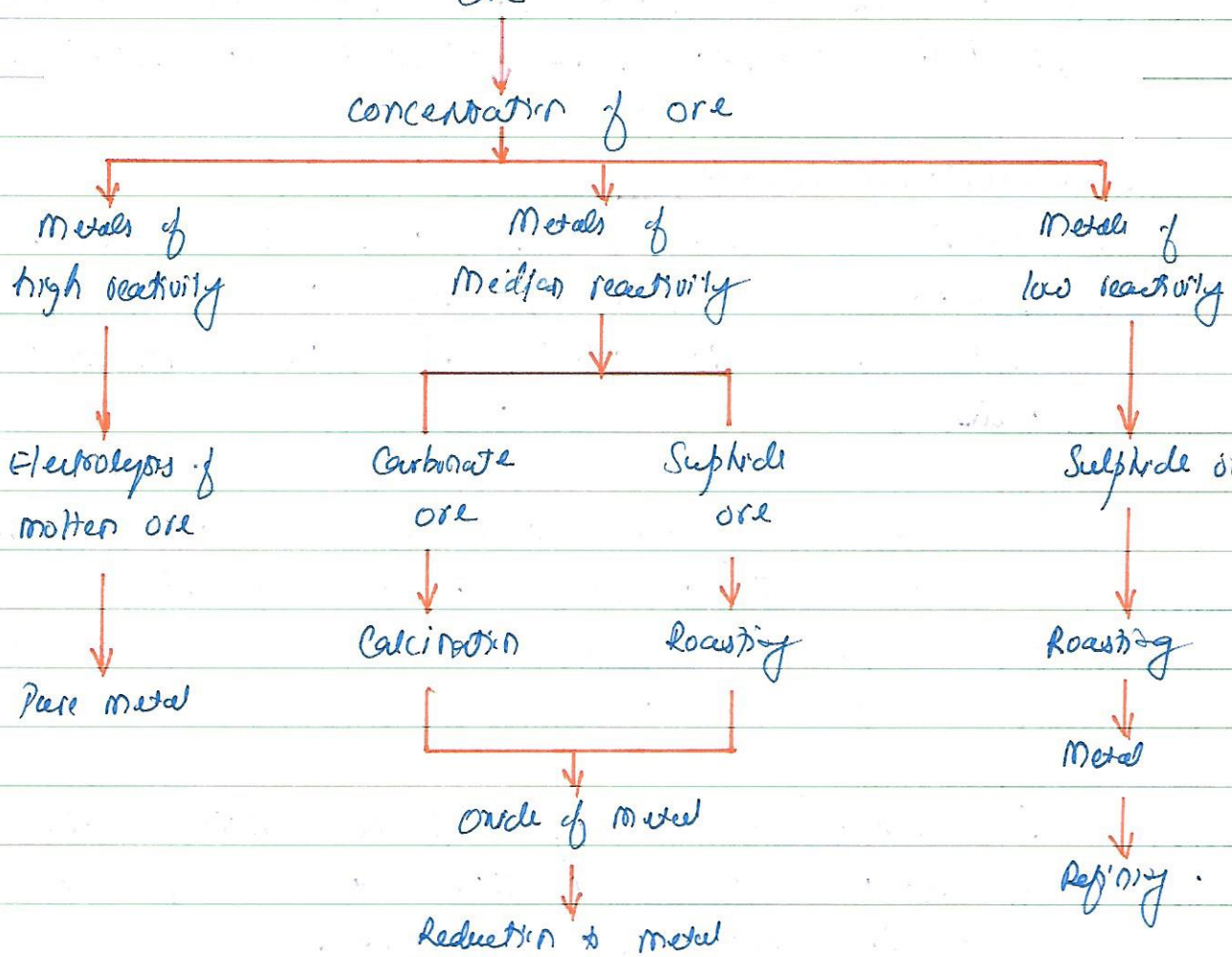
- The metals at the bottom of the activity series are the least reactive. They are often found in a free state. For example, gold, silver, platinum and copper are found in the free state.
- The metals at the top of the activity series (K, Na, Ca Mg and Al) are so reactive that they are never found in nature as free elements.
- The metals in the middle of the reactivity series (Zn, Fe, Pb etc) are moderately reactive. They are found in the earth's crust mainly as oxides, sulphides & carbonates.

Ores of many metals are oxides. This is because oxygen is a very reactive element and is very abundant on the earth.

Thus on the basis of reactivity, we can group the metals into the following three categories:

(i) Metals of low reactivity
(ii) Metals of medium reactivity
(iii) Metals of high reactivity

Steps involved in the extraction of metals from ores.



Concentration of ore (Enrichment of ore)

The unwanted impurities like sand, rocky material, earthy particles, limestone, mica etc present in an ore are called gangue.
The method used for removing gangue from ore depend on some differences in the physical properties or chemical properties of the ore and gangue.
Different separation techniques are accordingly employed.

Extraction Metals low in the Activity series.

Metal low in the activity series are very unreactive.
The oxides of these metals can be reduced to metal by heating alone.

(i) Cinnabar (HgS) is an ore of mercury. When heated in air, it first converted into mercury oxide (HgO). Mercury oxide is then reduced to mercury on further heating.

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

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

(ii) Copper which is found as Cu_2S in nature can be obtained from its ore by just heating in air.

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

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

Extraction of Metals in the Middle of the Activity Series.

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

Prior to reduction, the metal sulphides and carbonates must be converted into metal oxides.

Roasting:

Roasting is the process in which a sulphide ore is strongly heated in the presence of air to convert it into metal oxide.

$$2ZnS(s) + 3O_2(g) \xrightarrow{\text{Roasting}} 2ZnO(s) + 2SO_2(g)$$

Zinc Sulphide (Zinc blende ore) + Oxygen (from air) → Zinc Oxide + Sulphur dioxide.

Thus roasting converts zinc sulphide into zinc oxides.

Calcination:

Calcination is the process in which a carbonate ore is heated strongly in the absence of air to convert it into metal oxide.

$$ZnCO_3(s) \xrightarrow{\text{Calcination}} ZnO(s) + CO_2(g)$$

Zinc Carbonate (Calamine ore) → Zinc Oxide + Carbon dioxide.

Thus, calcination converts zinc carbonates into zinc oxide.

The metal ovides (obtained by calcination or roasting of ores) are converted to the free metal by using reducing agents like carbon, aluminium, sodium or calcium.

(i) Reduction of Metal ovides with Carbon :

In the reduction by carbon, the metal ovide is mixed with carbon (in the form of coke) and heated in a furnace. Carbon reduces the metal ovide to free metal.

Zinc metal is extracted by the reduction of its ovide with carbon (or coke). Thus, when zinc ovide is heated with carbon, zinc metal is produced:

$ZnO(s) + C(s) \longrightarrow Zn(s) + CO(g)$

(Zinc ovide) Carbon (Reducing agent) Zinc Metal Carbon monoxide.

Carbon is a cheap reducing agent but it contaminates the metal.

Iron metal is extracted from its ovide ore 'haematite' (Fe_2O_3) by reduction with carbon (in the form of coke). Tin and lead metals are also extracted by the reduction of their ovides with carbon. Even the less reactive metal copper is extracted by the reduction of its ovide with carbon.

(ii) Reduction of Metal ovide with aluminium : Mn and Cr metals are extracted by the reduction of their ovides with aluminium powder.

$3MnO_2(s) + 4Al(s) \longrightarrow 3Mn(l) + 2Al_2O_3 + Heat$

(Manganese dioxide) Aluminium Powder (Reducing Agent) Manganese Metal Aluminium ovide.

The highly reactive metals such as sodium, calcium, aluminium etc. are used as reducing agent because they can displace metals of lower reactivity from their compounds.

[These displacement reactions are highly exothermic. The amount of heat evolved is so large that the metals are produced in the molten state. The reaction of iron (III) oxide (Fe_2O_3) with aluminium is used to join railway tracks or cracked machine parts. This reaction is known as the Thermit reaction.]

$$Fe_2O_3(s) + 2Al(s) \longrightarrow 2Fe(l) + Al_2O_3(s) + Heat$$

Extraction of Highly Reactive Metals.

The highly reactive metals (K, Na, Ca, Mg and Al) are extracted by the electrolysis of their molten chlorides or oxides.

During electrolysis, the negatively charged electrode (cathode) acts as a powerful reducing agent by supplying electrons to reduce the metal ions into metal. The metals are always produced at the cathode (negative electrode). Chlorine is liberated at the anode (the positively charged electrode).

Electrolysis of molten sodium chloride. (NaCl)

At cathode: $2Na^+ + 2e^- \longrightarrow 2Na$

At anode: $2Cl^- \longrightarrow Cl_2 + 2e^-$

$2NaCl(l) \xrightarrow{\text{electrolysis}} 2Na(s) + Cl_2(g)$

The metal extracted by electrolysis method are very pure.

Similarly, aluminium is obtained by the electrolytic reduction of aluminium oxide (Al_2O_3)

Extraction of Sodium Metal

When electric current is passed through molten sodium chloride, it decomposes to form metal and chlorine gas:

$$2NaCl(l) \xrightarrow{\text{Electrolysis}} 2Na(s) + Cl_2(g)$$
(molten)

Molten NaCl contains free Na^+ and Cl^-.
The positive sodium (Na^+) are attracted to the cathode (negative electrode). The Na^+ take electrons from cathode and gets reduced to form sodium (atoms) (or sodium metal)

$$2Na^+ + 2e^- \longrightarrow 2Na \quad (\text{At cathode})$$

Thus, sodium metal is produced at the cathode (-ve electrode)

The Cl^- (negative chlorine ions) are attracted to the anode (positive electrode). The chlorine ions give electrons to the anode and get oxidised to form chlorine gas:

$$2Cl^- - 2e^- \longrightarrow Cl_2 \quad (\text{at anode})$$

Thus, chlorine gas is formed at the anode (+ve electrode)

Oxidation: is a chemical process in which there is loss of e^-
$Na \longrightarrow Na^+ + e^-$; $2Cl^- \longrightarrow Cl_2 + e^-$; $Fe^{2+} \longrightarrow Fe^{3+} + e^-$

Reducing agent: The substance whose loses electrons.

Reduction: In chemical process in which there is gain of e^-
$Cl_2 + 2e^- \longrightarrow 2Cl^-$; $S + 2e^- \longrightarrow S^{2-}$; $Fe^{3+} + e^- \longrightarrow Fe^{2+}$.

Oxidising agent: is the species which accepts electrons.

Refining of Metals

The process of purifying impure metals is called refining of metals. The most important and most widely used method for refining impure metals is electrolytic refining.

Electrolytic Refining :

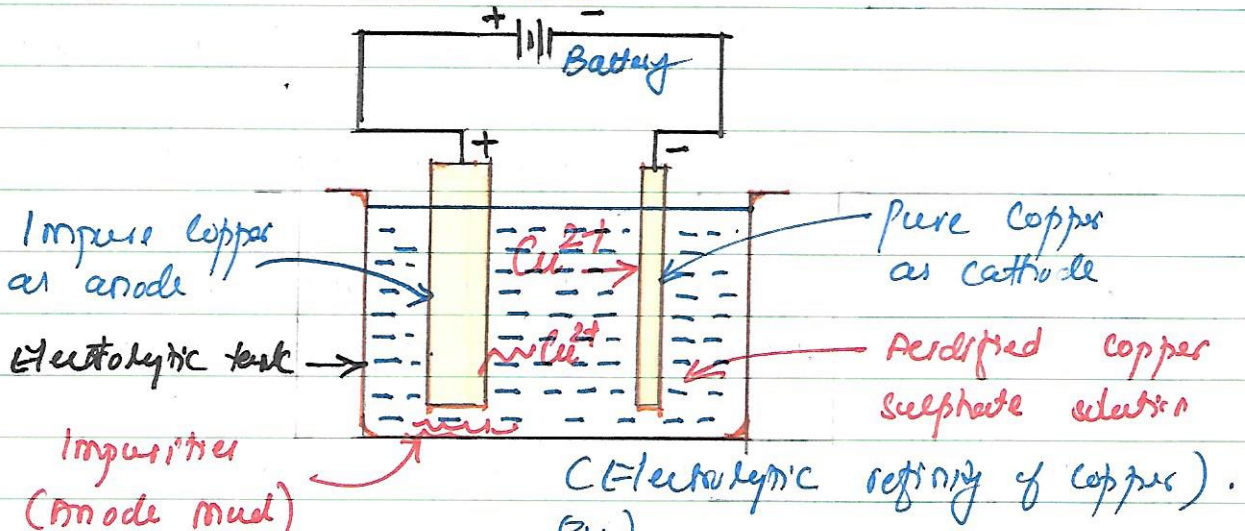
Many metals such as copper, zinc, tin, nickel, silver, gold etc., are refined electrolytically.

(a) A thick block of the impure metal is made anode. (It is always connected to the positive terminal of the battery).

(b) A thin strip of the pure metal is made cathode. (It is connected to the negative terminal of the battery).

(c) A water soluble salt (of the metal to be refined) is taken as electrolyte.

On passing the current through the electrolyte, the pure metal from the anode dissolve into electrolyte An equivalent amount of pure metal from the electrolyte is deposited on the cathode. The soluble impurities go into the solution, whereas insoluble impurities settle down at the bottom of the anode and are known as anode mud.



(Electrolytic refining of copper).

Corrosion

The eating up of metals by the action of air, moisture or a chemical (such as an acid) on their surface is called corrosion. Most of the metals corrode when they are kept exposed to damp air (or moist air).

- Silver articles becomes black after some time when exposed to air. This is because it reacts with sulphur in the air to form a coating of silver sulphide.

- Copper reacts with moist carbon dioxide in the air and slowly loses its shiny brown surface and gains a green coat. This green substance is copper carbonate.

- Iron when exposed to moist air for a long time acquires a coating of a brown flaky substance called rust. The corrosion of iron is called rusting.

Prevention of Rusting

1) The rusting of iron can be prevented by painting, oiling, greasing, galvanising, chrome plating, anodising or making alloys.
2) Galvanisation is a method of protecting steel and iron from rusting by coating them with a thin layer of zinc.
3) When iron is mixed with nickel and chromium, we get stainless steel, which is hard and does not rust.

An alloy is a homogeneous mixture of two or more metals, or a metal and a non-metal. The electrical conductivity and melting point of an alloy is less than that of pure metals.

Questions

(1) Define the following term (i) mineral (ii) ore (iii) gangue.

Ans) (i) Mineral : The element or compounds, which occur naturally in the earth's crust, are known as minerals.

(ii) Ore : If mineral contain very high percentage of a particular metal can be profitably extracted from it. These minerals are called ores.

(iii) Gangue : Impurities present in ores.

(2) Name two metal which are found in nature free state.

Ans) Gold, Silver and platinum.

(3) What chemical process is used for obtaining a metal from its oxide?

Ans) Reduction is used to obtain a metal from its oxide. Metal oxides are reduced by using suitable reducing agents such as carbon or by highly reactive metals to displace the metal from their oxide.
For example, zinc oxide is reduced to metallic zinc by heating with carbon.

$$ZnO(s) + C(s) \xrightarrow{\Delta} Zn(s) + CO(g)$$

Manganese dioxide is reduced to meganese (Mn) by treating it will aluminium powder. Aluminium displaces manganese from its oxide.

$$3MnO_2(s) + 4Al(s) \longrightarrow 3Mn(l) + 2Al_2O_3(s) + \Delta(Heat)$$

Oxide of more reactive metals are reduced by electrolysis.

(4) Metal oxides of zinc, magnesium and copper were heated with the following metals. In which cases will you find displacement reaction taking place?

	Zinc	Magnesium	Copper
Zinc oxide	No reaction	Displacement	No reaction
Magnesium oxide	No reaction	No reaction	No reaction
Copper oxide	Displacement	Displacement	No reaction.

(5) Which metal do not corrode easily?

Ans) Gold do not corrode easily because it is less reactive metal.

(6) What are alloys?

Ans) Alloys are homogeneous mixtures of two or more elements. The elements could be two metals, or a non-metal and metal. An alloy is formed by first melting the metal and then dissolving the other element in it. For example, steel is an alloy of iron and carbon.

(7) In the electrolytic refining of a metal M, what would you take as the anode, the cathode and the electrolyte?

Ans) Anode – Impure metal M

Cathode – Thin strip of pure metal M.

(8) Pradyush took sulphur powder on a spatula and heated it. He collected the gas evolved by inverting a test tube over it. (a) What will the action of gas on (i) dry litmus paper? (ii) moist litmus paper?
(b) Write balanced chemical equation.

Ans) (a) (i) There will be no action on dry litmus paper

(ii) moist blue litmus paper to red.

(b) $S(s) + O_2(g) \longrightarrow SO_2(g)$, $SO_2 + H_2O(l) \longrightarrow H_2SO_4(aq)$

B7) Sulphuric acid.

(9) State two ways to prevent the rusting of iron.

Ans) Refer to notes.

(10) What type of oxides are formed when non-metals combine with oxygen?

Ans) Non-metal combine with oxygen & form acidic oxide

$S(s) + O_2(g) \longrightarrow SO_2(g)$ (acidic in nature)

(11) Explained why copper vessels are cleaned with lemon or tamarind.

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 basic copper carbonate and dissolves the layer. Therefore, tarnished copper vessels are cleaned with lemon or tamarind juice to give the surface of copper vessel its characteristic lustre.

(12) Give reasons why copper is used to make hot water tanks and not steel.

Ans) Copper does not react with cold water, hot water or steam. However, iron reacts with steam.

$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$

(Steam) — Iron (II,III) oxide