

Student Learning Outcomes

After studying this chapter, students will be able to:

- Define redox reactions as simultaneous oxidation and reduction in terms of oxygen, hydrogen, electrons and changes in oxidation state
- Use roman numerals to indicate oxidation number of an element in a compound
- Identify oxidizing and reducing agents in a redox reaction
- Recognize that the oxidation number of elements in their free state is zero
- Derive the formula of ionic compounds from ionic charges and oxidation numbers
- Identify that the oxidation number of a monoatomic ion is the same as the charge on the ion
- Explain that the sum of the oxidation numbers in a neutral compound is zero
- Explain that the sum of the oxidation numbers in an ion is equal to the charge on the ion
- Identify redox reactions by the colour changes involved when using acidified aqueous potassium manganate (VII) to (II) or aqueous potassium iodide
- Define electrolysis as decomposition of ionic compound, in molten or aqueous solution, by passage of electric current
- Identify and label in simple electrolytic cells, the anode (+), cathode (-), electrolyte and direction of flow of electrons in external circuit,
- Describe the transfer of charge in external circuit, movement of ions in the electrolyte and transfer of electrons at electrodes
- Identify the products formed at electrodes and describe the observations made during the electrolysis of molten lead (II) chloride, concentrated dilute sulfuric acid using inert electrodes (platinum or carbon/graphite)
- State that hydrogen- oxygen fuel cell uses hydrogen and oxygen to produce electricity with water as the only chemical product
- Describe the advantages and disadvantages of using hydrogen-oxygen fuel cells in comparison with gasoline /petrol engines in vehicles.
- Define corrosion and discuss methods to prevent it. (some examples may include barrier method such as using paint, galvanizing, electroplating; sacrificial protection such as using magnesium blocks in ships)
- Identify the products formed at electrodes and describe the observations made during the electrolysis of dilute copper (II) sulphate using inert electrode or copper electrode
- Predict the identity of products of electrolysis of a halide compound in dilute or concentrated solution
- Construct ionic half- equations for reaction at either electrode.
- Describe electroplating and its applications.
- Sketch a schematic diagram for a voltaic cell e.g., Daniel cell
- Use the voltage data given for voltaic cells to determine order of reactivity of any two metals

Subjective Type Questions

Long Questions prepared in the light of new Examination Techniques
(Knowledge, Understanding, Application, Analytical & Conceptual)

16.1 Oxidation Number of Atoms in Compounds and Ions

16.2 Writing Chemical Formulae of Ionic Compounds by applying Oxidation Numbers

16.3 Writing Chemical formulae of Ionic compounds using ionic charges

16.4 Oxidation and Reduction

16.5 Oxidation - Reduction Reactions

Q.1: How the oxidation number of atoms in compounds and ions can be found out?

Ans. Oxidation Number

The oxidation number is an apparent charge on an atom in a compound or an ion. It may have a positive or a negative sign.

Oxidation Number of Atoms in Compounds and Ions

The oxidation number of atoms can be found out by applying the following rules.

1. Atom present in a free element has an oxidation number zero.

Example: The oxidation number of sodium in sodium metal is zero. Similarly, oxidation number of oxygen in ozone(O_3) is also zero.

2. For monoatomic ions, the oxidation number is equal to the charge present on the ion.

Example: The oxidation number of Ca^{2+} is +2 while that of Al^{3+} is +3.

3. The elements present in group 1 (except hydrogen) in the periodic table are assigned oxidation number +1, those of group 2, +2 and those of group 3, +3.

4. In binary compounds the more electronegative atom is always assigned a negative oxidation number. For example, fluorine, being the most electronegative atom, will always be given -1 oxidation number. Similarly, in NH_3 and H_2S , nitrogen and sulphur are assigned the oxidation number of -3 and -2 respectively.

5. The elements of group 17 are normally given oxidation number -1 when they are combined with less electronegative element.

Example: Chlorine in all such compounds as HCl , PCl_3 , $CHCl_3$, has been assigned an oxidation number -1.

6. Hydrogen is assigned an oxidation number +1 when it is in combination with more electronegative atoms while it is given oxidation number -1 when it is combined with less electronegative atoms.

Example: In HCl , the oxidation number of hydrogen is +1, while in sodium hydride

(NaH) it bears an oxidation number -1.

- The element oxygen is assigned an oxidation number -2 in most of its compounds.
- The sum of oxidation numbers of all the atoms present in a neutral compound is always zero. In a polyatomic ion the sum of oxidation numbers is equal to the charge present on that ion.

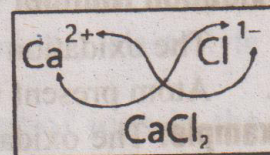
Q.2: Explain with examples how the formula of ionic compound can be written using oxidation numbers.

Ans. Writing Chemical Formulas of Ionic Compounds by applying Oxidation Numbers:

We can write the chemical formula of a compound if we know the oxidation numbers of its constituent elements. The periodic table and the electronic structure of the atoms can help us in predicting the principal oxidation number of the element. The formula of the compound may then be written using the oxidation number of the atoms present in that compound. For example, we shall write the formula of calcium chloride in the following way.

The element calcium is present in group II of the periodic table and its principal oxidation number is +2.

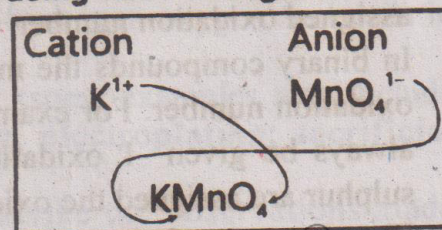
Chlorine is present in Group 17 and we know that when it is attached with less electronegative element in binary compounds its oxidation number is -1. The formula for calcium chloride will, therefore, be CaCl_2 because only with this formula the sum of the oxidation numbers of both the elements will add up to zero.



Q.3: Discuss with examples how the chemical formulas of ionic compounds can be written using ionic charges.

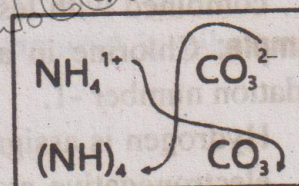
Ans. Writing Chemical Formulas of Ionic Compounds using Ionic Charges:

In order to write the formula of potassium permanganate, write first the formulae of permanganate ion, MnO_4^{1-} and potassium ion K^{1+} . Then combine the two ions together to form an electrically neutral compound.



The formula for potassium permanganate will, therefore, be KMnO_4 .

When writing the formulae of ionic compounds containing more than one polyatomic ions, we take the formula of the ion in brackets and indicate the number of such ions with a subscript. The formula for ammonium carbonate will be written as follows.



The formula for ammonium carbonate will, therefore, be $(\text{NH}_4)_2\text{CO}_3$.

Q.4: Write a note on oxidation.

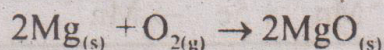
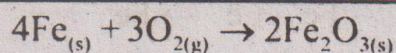
Ans. Oxidation

Oxidation is described in three principal ways.

1. Oxidation is defined as a process in which oxygen is either added or hydrogen is removed.

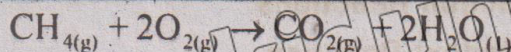
For example;

- (i) Oxygen combines with other elements in the formation of their oxides. Elements in these processes are said to be oxidized and oxygen is reduced. Rusting of iron or burning of magnesium in air are oxidation processes.



In the above reactions Fe and Mg are being oxidized while oxygen is being reduced. Oxygen is an oxidising agent while iron and magnesium are reducing agents.

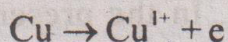
- (ii) Hydrogen is removed from a compound by oxygen in the following examples; hydrogen is removed from CH_4 and HCl by oxygen or a compound containing oxygen.



CH_4 and HCl in these reactions are being oxidized while O_2 and MnO_2 are being reduced. CH_4 and HCl are called reducing agents while O_2 and MnO_2 are called oxidising agents.

2. Oxidation is a process in which electron or electrons are lost.

For example; Copper may lose one or two electrons to give Cu^{1+} or Cu^{2+} ions respectively. In both these processes, we say, copper is being oxidised. Hence, copper is called an reducing agent.



3. Oxidation is process in which oxidation number or oxidation state of an element increases.

In the above two processes, the oxidation number of copper has increased from zero to +1 [$\text{Cu}(\text{I})$] and +2 [$\text{Cu}(\text{II})$] respectively.

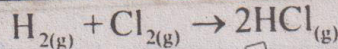
Q.5: Write a note on reduction.

Ans. Reduction

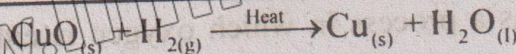
Reduction is described in three principal ways.

1. Reduction is defined as either addition of hydrogen or removal of oxygen. For example,

- (i) Hydrogen is added in chlorine to give hydrogen chloride gas.

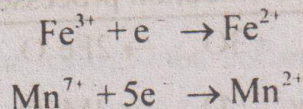


- (ii) In the following reaction, hydrogen is heated with metal oxide when oxygen comes out of the compound and combines with hydrogen to give a molecule of water.



In the above examples, Cl_2 and CuO are being reduced and hydrogen (in both cases) is being oxidized. Cl_2 and CuO are called oxidizing agents while hydrogen in both the reactions will be reducing agent.

2. Reduction is a process in which electron or electrons are gained. For example,



In both the above examples, ions of iron and manganese are being reduced. Hence, they are called oxidising agents.

3. Reduction is a process in which oxidation number of an element decreases. In the above examples oxidation number of iron is reduced from +3 [Fe(III)] to +2 [Fe(II)] while oxidation number of manganese is reduced from +7 [Mn(VII)] to +2 [Mn(II)].

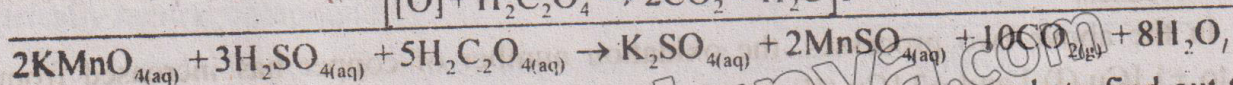
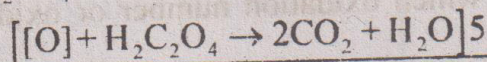
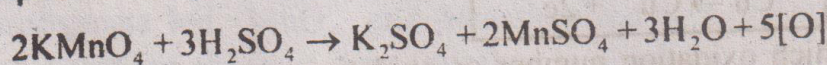
Q.6: Explain Oxidation-Reduction Reactions.

Ans. Oxidation-Reduction (Redox) Reaction

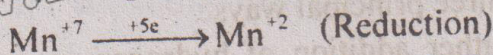
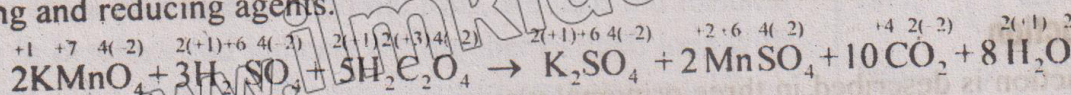
A chemical reaction which involves simultaneous oxidation and reduction is called an oxidation-reduction or redox reaction.

Principle: The reaction may involve the complete transference of electron or electrons to form ionic bonds or a partial transfer or shift of electron or electrons to form covalent bonds.

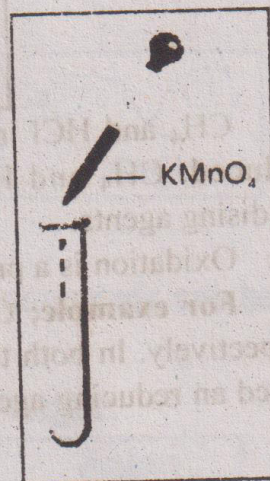
Explanation: Potassium permanganate acts as a very powerful oxidising agent in acidic medium. It oxidizes oxalic acid into carbon dioxide. During the reaction the dark purple solution of potassium permanganate changes to colourless. In the presence of dilute sulphuric acid the reaction proceeds as follows.



Let us assign oxidation number to each element in these compounds to find out the oxidising and reducing agents.



Oxidation number of manganese has decreased from +7 to +2 and that of carbon

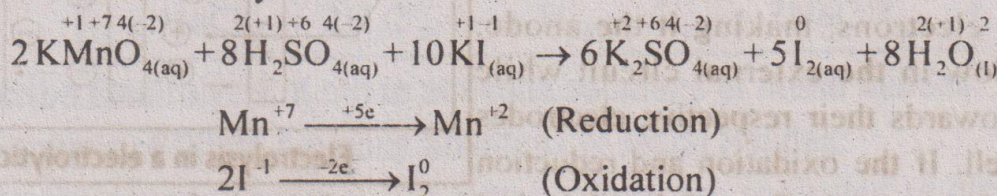


Decolourization

has increased from +3 to +4.

So, each manganese atom accepts five electrons to get reduced and each carbon atom loses one electron to get oxidized. KMnO_4 is, therefore, an oxidizing agent while oxalic acid in this reaction acts as a reducing agent.

Aqueous potassium iodide is also oxidized to iodine in acidified potassium permanganate. The formation of iodine in this reaction results in a colour change, typically from colourless to yellow-brown.



In this reaction, the permanganate ion (MnO_4^-) is reduced to manganese (II) ion (Mn^{2+}) and the iodide ion (I^-) is oxidized to iodine (I_2). Potassium permanganate, therefore, acts as an oxidizing agent while potassium iodide acts as a reducing agent.

16.6	Electrolysis
16.7	Electrolysis of a Concentrated Aqueous Solution of Sodium Chloride (Brine) using inert electrodes (Graphite / Pt)
16.8	Electrolysis of Dilute and Concentrated Aqueous solutions of Metal Halides
16.9	Electrolysis of Concentrated Aqueous Solutions of Metal Halides

Q.7: Define electrolysis. How does the process of electrolysis take place?

Ans. Electrolysis:

The process in which an electricity is passed through fused ionic compounds or their aqueous solution, they are decomposed into their constituents, is called electrolysis.

Electrolytes: The compounds which are decomposed into their constituents during the electrolysis are called electrolytes.

Chemical Reaction: During electrolysis, a non-spontaneous chemical reaction takes place with the help of electrical energy.

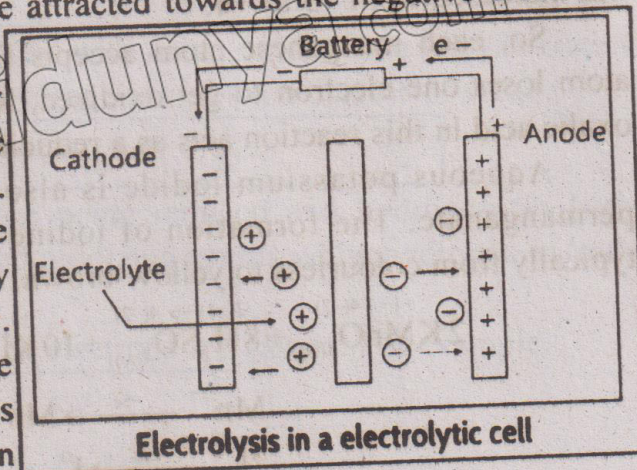
Electrolytic Cell: The process of electrolysis can be carried out in an electrolytic bath called an electrolytic cell with the help of two electrodes through which electric current enters and leaves the cell. These electrodes may be made of graphite or any other metal.

Anode: It is a positive electrode through which electrons enter the external circuit.

Cathode: It is a negative electrode through which electrons leave the external circuit.

Explanation: Electrolyte present in the cell is in the form of freely moving ions. During electrolysis, the electrolyte is decomposed into its components by applying an external voltage. Thus external voltage source (battery) forces the non-spontaneous reaction to occur by pumping electrons towards one of the electrodes, thereby, giving it a negative

charge. The positive ions of the electrolyte are attracted towards the negative electrode and get reduced by accepting electrons from it, making it the cathode. The battery simultaneously withdraws electrons from the other electrode, giving it a positive charge. The negative ions of the electrolyte are attracted towards this and get oxidized by losing their electrons, making it the anode. Electrons flow in the external circuit while ions move towards their respective electrodes inside the cell. If the oxidation and reduction at the electrodes cease, the flow of electricity in the external circuit also stops.

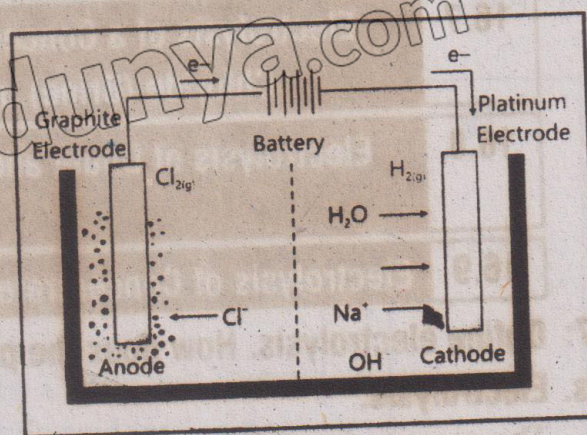


Q.8: Explain electrolysis of a concentrated aqueous solution of sodium chloride (Brine) using inert electrodes (Graphite/Pt).

Ans. Electrolysis of a Concentrated Aqueous Solution of Sodium Chloride (Brine) using inert electrodes (Graphite/Pt):

The electrolysis of a concentrated solution of sodium chloride, also called brine, produces hydrogen gas at the cathode, chlorine gas at the anode, and sodium hydroxide solution as a byproduct.

The electrolytic cell used in this electrolysis. The cathode used in this cell is made of platinum while the anode must be made of graphite to avoid the reaction with chlorine gas.



Electrolysis of brine

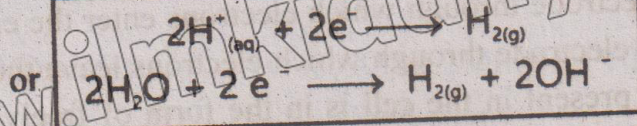
The following ions are present in the solution

$\text{Na}^+_{(\text{aq})}$, $\text{Cl}^-_{(\text{aq})}$ from sodium chloride

$\text{H}^+_{(\text{aq})}$, $\text{OH}^-_{(\text{aq})}$ from water

Na^+ and H_2O , both migrate towards the cathode. Since water molecules and hydrogen ions have a greater tendency to get reduced as compared to sodium ions (Na^+), so they pick up electrons from the cathode to form hydrogen gas.

At the Cathode:



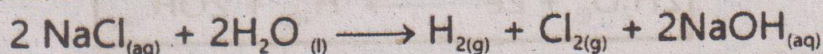
The ionic equilibrium of water is, thus, disturbed due to the discharge of H^+ and more water molecules are ionized to restore it. The solution around the cathode becomes alkaline due to the formation of hydroxide ions (OH^-).

At the Anode:



The solution gradually gets alkaline due to the presence of OH^- . The volumes of hydrogen and chlorine gases obtained are equal.

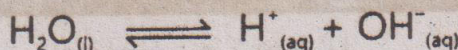
The overall equation for this electrolysis is:



Q.9: Describe electrolysis of dilute aqueous solutions of metal halides.

Ans. Electrolysis of Dilute Aqueous Solutions of Metal Halides:

When the aqueous solutions of metal halides are subjected to electrolysis, the situation becomes more complicated due to the presence of water. Water is a very weak electrolyte and upon ionization it gives very small amount of hydrogen and hydroxide ions.

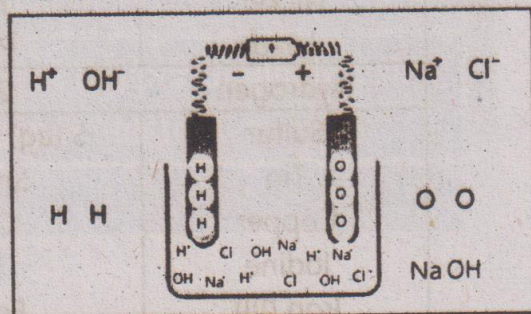


This will mean that the electrolyte will consist of more than one type of ions which will approach both cathode and anode. Hence there can be a choice over which ions will get discharged.

For example, if an aqueous solution of sodium chloride is electrolyzed, both sodium and hydrogen cations are attracted towards the cathode and both chlorides and hydroxide ions are attracted towards the anode. Which of these ions will preferentially be discharged depends upon the tendency of the elements to lose or gain electrons. The elements present at the top of electrochemical series possess more tendency to lose electrons as compared to the elements present at the bottom.

If a metal is below hydrogen in the electrochemical series, then the metal ion is discharged. Metals like this include copper and silver. If dilute copper chloride is electrolyzed, copper metal is obtained at the cathode.

If a metal is placed higher in the electrochemical series then hydrogen ion is discharged. Metals like this include lithium, sodium and magnesium. If dilute solution of sodium chloride is electrolyzed, hydrogen gas is obtained at the cathode.

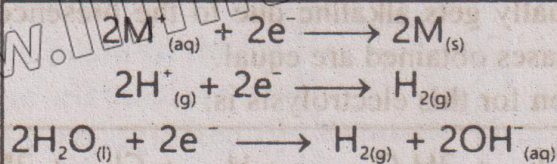


Electrolysis of dilute solutions of halides

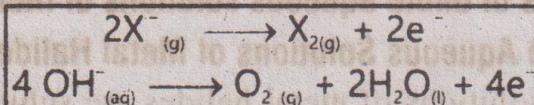
Metals like lead and nickel present a more complicated picture. If aqueous solutions of halides of these metals are electrolyzed, then the products obtained at the cathode depends on the concentrations of the halide solutions. If the solution is concentrated then metal is deposited at the cathode. If the solution is very dilute, hydrogen is evolved. At in-between concentrations both metal and hydrogen are obtained.

At anode, the electrolysis of dilute solutions of halides will deposit oxygen because hydroxide ions are preferentially discharged due to their higher concentrations.

Reaction at the Cathode:



Reaction at the Anode:



Electrochemical Series

Stronger reducing agent	Electrode	Half Reaction	Standard Potential (V)	Weaker oxidizing agent
	Lithium	$Li^{+}(aq) + e^{-} \rightleftharpoons Li(s)$	-3.04	
	Potassium	$K^{+}(aq) + e^{-} \rightleftharpoons K(s)$	-2.93	
	Calcium	$Ca^{2+}(aq) + 2e^{-} \rightleftharpoons Ca(s)$	-2.86	
	Sodium	$Na^{+}(aq) + e^{-} \rightleftharpoons Na(s)$	-2.71	
	Magnesium	$Mg^{2+}(aq) + 2e^{-} \rightleftharpoons Mg(s)$	-2.37	
	Aluminum	$Al^{3+}(aq) + 3e^{-} \rightleftharpoons Al(s)$	-1.66	
	Water	$2H_2O(l) + 2e^{-} \rightleftharpoons H_2(g) + 2OH^{-}(aq)$	-0.82	
	Zinc	$Zn^{2+}(aq) + 2e^{-} \rightleftharpoons Zn(s)$	-0.76	
	Iron (II)	$Fe^{2+}(aq) + 2e^{-} \rightleftharpoons Fe(s)$	-0.44	
	Cadmium	$Cd^{2+}(aq) + 2e^{-} \rightleftharpoons Cd(s)$	-0.40	
	Nickel	$Ni^{2+}(aq) + 2e^{-} \rightleftharpoons Ni(s)$	-0.25	
	Lead	$Pb^{2+}(aq) + 2e^{-} \rightleftharpoons Pb(s)$	-0.12	
	Hydrogen	$2H^{+}(aq) + 2e^{-} \rightleftharpoons H_2(g)$	0	
	Sulfur	$S(aq) + 2H^{+}(aq) + 2e^{-} \rightleftharpoons H_2S(aq)$	0.14	
	Tin	$Sn^{4+}(aq) + 2e^{-} \rightleftharpoons Sn^{2+}(aq)$	0.15	
	Copper	$Cu^{2+}(aq) + 2e^{-} \rightleftharpoons Cu(s)$	0.34	
	Iodine	$I_2(g) + 2e^{-} \rightleftharpoons 2I^{-}(aq)$	0.53	
	Iron (III)	$Fe^{3+}(aq) + e^{-} \rightleftharpoons Fe^{2+}(aq)$	0.77	
	Silver	$Ag^{+}(aq) + e^{-} \rightleftharpoons Ag(s)$	0.79	
	Bromine	$Br_2(g) + 2e^{-} \rightleftharpoons 2Br^{-}(aq)$	1.06	
	Chlorine	$Cl_2(g) + 2e^{-} \rightleftharpoons 2Cl^{-}(aq)$	1.35	
	Gold	$Au^{+}(aq) + e^{-} \rightleftharpoons Au(s)$	1.69	
	Hydrogen peroxide	$H_2O_2(aq) + 2H^{+}(aq) + 2e^{-} \rightleftharpoons 2H_2O(l)$	1.77	
	Fluorine	$F_2(g) + 2e^{-} \rightleftharpoons 2F^{-}(aq)$	2.87	
Weaker reducing agent				Stronger oxidizing agent

Q.10: Discuss the electrolysis of concentrated solutions of metal halides.

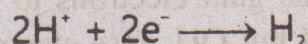
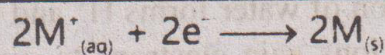
Ans. Electrolysis of Concentrated Aqueous Solutions of Metal Halides:

If the concentrated aqueous solutions of metal halides are electrolyzed the hydrogen ions are discharged at the cathode if the metal is present higher than hydrogen in electrochemical series. For example, electrolysis of concentrated aqueous solution of sodium chloride will produce hydrogen gas at the cathode.

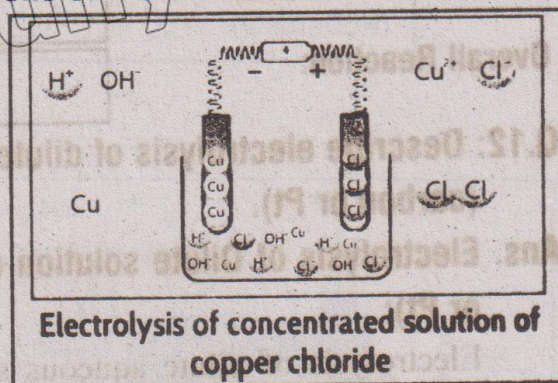
However, for less reactive metals like copper and silver the metal ions are discharged at the cathode.

At the anode, the halide ions are preferentially discharged due to their higher concentration.

Reaction at the Cathode:



Reaction at the Anode:



16.10

Electrolysis of Molten Lead (II) chloride

16.11

Electrolysis of Dilute solution of Sulphuric acid using inert electrodes (carbon or Pt)

16.12

Electrolysis of Dilute Aqueous Solution of Copper Sulphate using inert electrodes or Copper electrodes

16.13

Fuel Cells

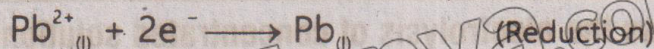
Q.11: Explain electrolysis of molten lead (II) chloride.

Ans. Electrolysis of Molten Lead (II) Chloride:

Lead (II) chloride is solid at room temperature. To make it an electrolyte it must be heated upto 501°C to get molten lead chloride. Heat will render ions free to move and conduct electricity. The molten lead is then taken in a heated crucible where its temperature is maintained above 500°C.

Two inert electrodes (graphite or platinum) are then immersed in it. Upon electrolysis the lead (II) ions will move towards the cathode where they are discharged forming lead in the liquid state. The chloride ions migrate towards the anode where they lose electrons to give chlorine gas. The electrolysis must be carried out at the temperature above the melting point of lead chloride.

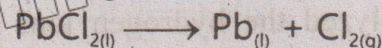
Reaction at the Cathode:



Reaction at the Anode:



Overall Reaction:



Q.12: Describe electrolysis of dilute solution of sulphuric acid using inert electrodes (carbon or Pt).

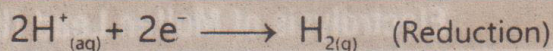
Ans. Electrolysis of Dilute solution of Sulphuric Acid using inert electrodes (Carbon or Pt):

Electrolysis of dilute aqueous solution of sulphuric acid produces hydrogen and oxygen gases. In other words this electrolysis involves the decomposition of water. Inert electrodes like platinum or graphite are used for this electrolysis because they did not react with the electrolyte.

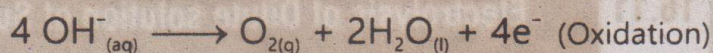
In an aqueous solution of dilute sulphuric acid H^{+} and SO_4^{2-} are given by sulphuric acid while the ionization of water forms H^{+} and OH^{-} .

At the cathode, hydrogen ions gain electrons to give hydrogen gas. At the anode both SO_4^{2-} and OH^{-} are attracted but OH^{-} lose electrons to give oxygen gas. Sulphate ions are not discharged at the anode because they are difficult to be oxidized compared to hydroxide ions. As the electrolysis moves ahead, the concentration of sulphuric acid increases because the water present is being decomposed.

Reaction at the Cathode:



Reaction at the Anode:

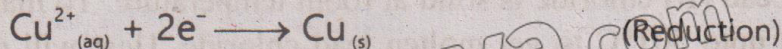


Q.13: Discuss electrolysis of dilute aqueous solution of copper sulphate using inert electrodes or copper electrodes.

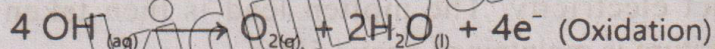
Ans. Electrolysis of Dilute Aqueous solution of Copper Sulphate using Inert electrodes or Copper electrodes:

The electrolysis of aqueous solution of copper(II) sulphate with inert electrodes result in the decomposition of water and the copper(II) sulphate. Copper metal is deposited at the cathode (inert) and oxygen gas is produced at the anode (inert).

Reaction at the Cathode:



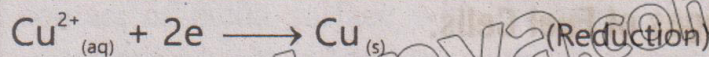
Reaction at the Anode:



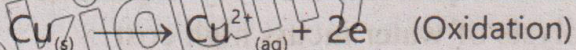
Since the concentration of copper ions decreases slowly so the blue colour of the solution fades away with the passage of time.

If copper electrodes are used instead of inert electrodes, the copper metal atoms from the copper anode start going into the solution in the form of copper ions and the electrode mass starts decreasing.

Reaction at the Cathode:



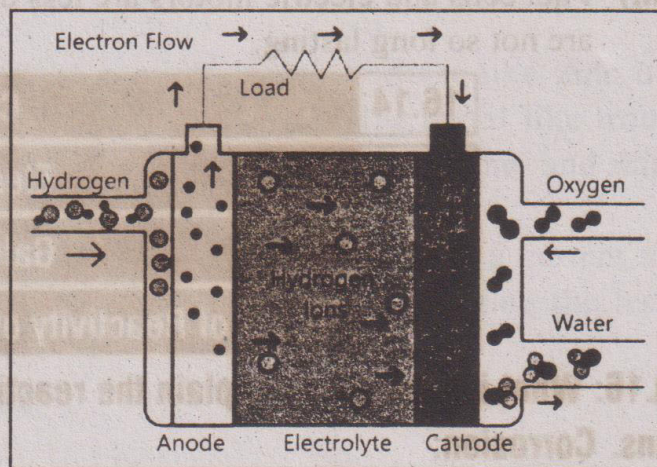
Reaction at the Anode:



Q.14: Describe the working of a fuel cell.

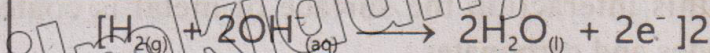
Ans. Working of a Fuel Cell:

Fuel cells provide a method in which chemical energy is converted to electrical energy. These cells use gaseous fuels like hydrogen and oxygen to produce electricity. A fuel cell is composed of two hollow tubes which act as electrodes. They are made of porous compressed carbon saturated with platinum which acts as a catalyst. The electrolyte used in this cell is aqueous potassium hydroxide. At the anode, hydrogen is oxidized to water and at the cathode oxygen is reduced to hydroxide ions.

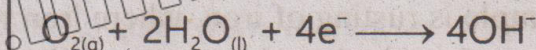


Fuel cell

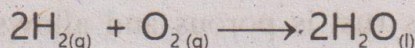
At anode:



At Cathode:



Overall Reaction:



The fuel cell produces electricity and water continuously provided the reactants are continuously supplied. They are very efficient because they convert 75% of the bond energy of fuels to electricity.

Q.15: Discuss advantages and disadvantages of fuel cells.

Ans. Advantages of Fuel Cells:

- (i) The hydrogen-oxygen fuel cells have a clear advantage over the traditional fossil fuels, petrol and diesel, used to drive the vehicles. They produce zero emission of carbon dioxide because water is the only product. Fossil fuels, on the other hand, produce carbon dioxide, nitrogen oxides and sulphur oxides which are very harmful to the environment.
- (ii) Fuel cells can convert a large proportion of chemical energy present in fuels into electrical energy as compared to the energy conversion of the fossil fuels. In other words, they are more efficient.
- (iii) Fuel cells produce less noise pollution.
- (iv) Hydrogen used in fuel cells can be obtained from renewable sources like solar or wind power. This fact renders fuel cells more sustainable.

Disadvantages of Fuel Cells:

- (i) Fuel cells at present, are very expensive because the production of hydrogen by electrolysis of water incurs high cost.
- (ii) Hydrogen being highly inflammable requires very special storage and transportation facilities which further increases the cost of fuel cells.
- (iii) Fuel cells and electric motors are less durable than petrol and diesel engines so they are not so long lasting.

16.14

Corrosion

16.15

Electroplating

16.16

Galvanic cells

16.17

Order of Reactivity of Metals using Voltage Data

Q.16: What is corrosion? Explain the reactions involved in corrosion of iron.

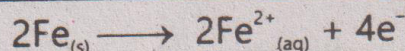
Ans. Corrosion:

Corrosion is a natural process due to which a metal suffers gradual decay when it interacts chemically with its environment.

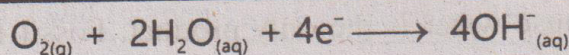
During this interaction the surface of metal is coated with compounds such as oxides, sulphides and carbonates.

Example: A common example is rusting of iron, where iron reacts with oxygen and water present in the atmosphere to form hydrated iron (III) oxide. Once iron begins to rust, the process will continue as rust is porous and allows both air and moisture to come into contact with fresh metal underneath. Hence, all the metal will corrode given enough time.

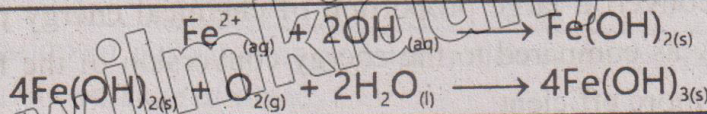
The process of corrosion of iron appears to require two important chemicals, oxygen and water, which are involved in the following electrochemical reactions. Iron loses electrons to form Fe (II) ions.



Oxygen gains these electrons in the presence of water to form hydroxide ions.



Fe^{2+} then react with more oxygen and water to form iron (III) hydroxide which is rust.



Q.17: How can corrosion be prevented?

Ans. Prevention of Corrosion:

Coating of the metal surface: Corrosion can be avoided by coating the metal surface with

some barriers that prevent the contact of metal surface with the atmosphere. For this purpose, the metal surfaces are protected by either greasing with oil or painting. Other methods of protection are galvanizing, electroplating and sacrificial protection.

Galvanization: Galvanization is a common method to protect iron or steel from rusting. It involves the dipping of an iron sheet in a bath of molten zinc. A layer of zinc metal gets deposited on the iron sheet and prevents rusting. Tin cans are actually steel cans which have been coated with a thin layer of tin.

Sacrificial Protection: In sacrificial protection a more reactive metal like zinc or magnesium is used which will corrode in place of the less reactive metal like iron. Magnesium is often employed to protect steel in buried fuel tanks and pipe lines and ship hulls submerged in sea water.

The difference in the activity of the two metals (Mg and Fe) causes a current to flow between them, producing corrosion on the more active metal protecting the less active metal iron.

Q.18: Explain electroplating giving examples.

Ans. Electroplating:

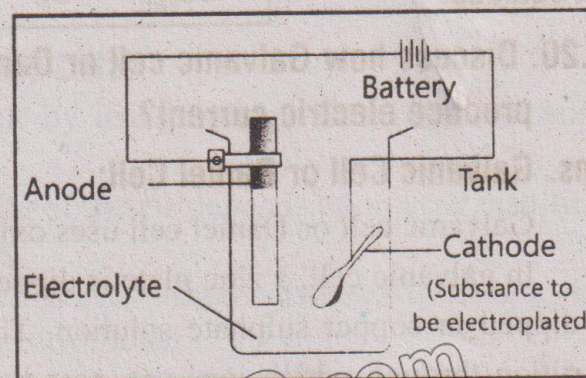
It is a process in which a metal is deposited on another metal electrolytically.

Objectives of Electroplating: There are three main objectives of electroplating.

- (i) **Decoration:** To deposit the noble metals like gold and silver on the inexpensive metals to enhance their beauty.
- (ii) **Protection:** To protect the metals from rusting as well as from the reaction of organic acids.
- (iii) **Repair:** To weld the broken parts of the machinery by electrodeposition of metals. Generally copper, silver, chromium, nickel and gold are electroplated.

Procedure of Electroplating:

The article to be electroplated is first cleaned with sand and then washed with caustic soda solution and finally with plenty of water. This article is then, made a cathode, while the metal to be deposited is made anode. The electrolyte is a salt of the metal being deposited and the electroplating is carried out in a tank made of cement, glass or wood. It is called an electrolytic tank.



The electrolyte chosen should have the following properties:

- (i) very soluble in water
- (ii) good conductor
- (iii) not easily oxidized or reduced or hydrolyzed
- (iv) inexpensive

Electrodes are dipped in the electrolytic tank containing the electrolyte and electric current is passed. A thin layer of the noble metal is deposited on the article to be electroplated. The following conditions should be fulfilled to get good deposits.

- i. High current density
- ii. Low temperature
- iii. Higher concentration of the metal in its electrolyte

Q.19: Write a note on the following:

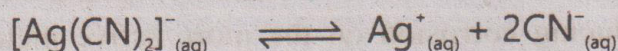
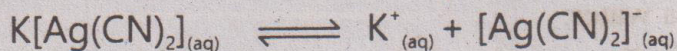
(a) Silver Plating

(b) Copper Plating

Ans. (a) Silver Plating:

The procedure for silver plating is the same as in electroplating except that the electrolyte used is the solution of potassium argento cyanide, $K[Ag(CN)_2]$, and the anode is of pure silver.

The following ions are formed in the electrolytic solution.

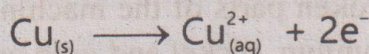


When the electric current is passed, Ag^+ ions migrate towards the cathode and deposit after picking up electrons. Cyanide ions move towards the anode, where they react with silver to form $AgCN$, which combines with KCN to reproduce the electrolyte, $K[Ag(CN)_2]$.

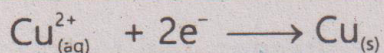
(b) Copper Plating:

In copper plating, the anode is made of copper metal and the electrolyte is $CuSO_4$ to which a few drops of sulphuric acid are added to prevent its hydrolysis. The article to be electroplated is made the cathode. The following electrolytic reactions take place, at the electrodes.

At Anode:



At cathode:



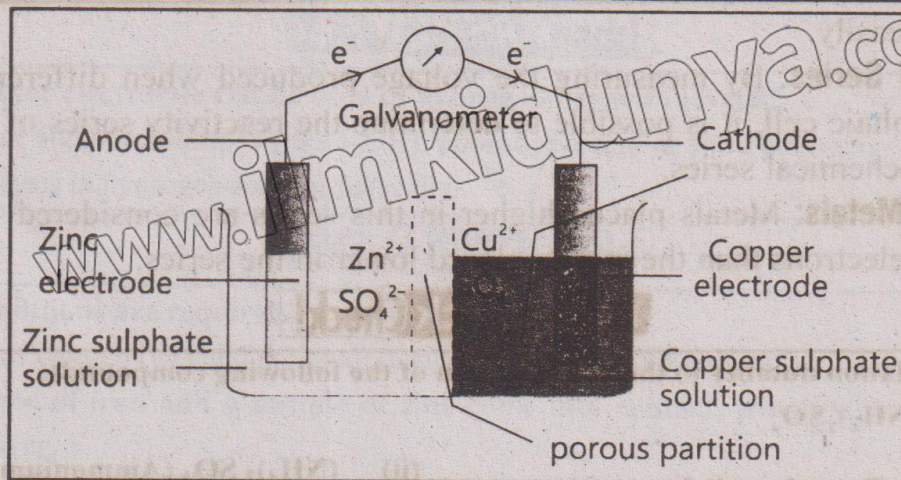
Q.20: Discuss how Galvanic cell or Daniel cell uses oxidation-reduction reactions to produce electric current?

Ans. Galvanic Cell or Daniel Cell:

Galvanic cell or Daniel cell uses oxidation reactions to produce electric current.

In galvanic cell, a zinc plate is dipped in zinc sulphate solution while a copper plate is dipped in copper sulphate solution. The two solutions are kept separate by a porous partition through which ions can pass but the solutions cannot. The two electrodes are connected by a wire through a galvanometer which measures the electric current.

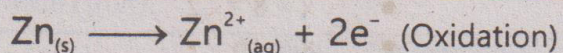
At the surface of zinc electrode, each zinc atom loses two electrons and goes into the solution as zinc ions. These electrons travel through the external wire to copper electrode where they are used to discharge copper ions into copper atoms which are deposited at this electrode. An oxidation reaction takes place at zinc electrode and it is called anode while reduction takes place at copper electrode and it is called cathode.



Daniel Cell

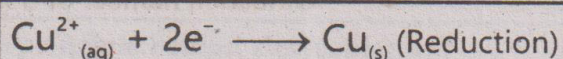
The following reactions take place at the two electrodes.

At anode:



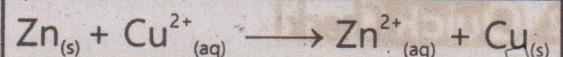
Standard Oxidation Potential = + 0.76

At cathode:



Standard Reduction Potential = + 0.34

Overall cell reaction:



Cell voltage $E^{\circ} = 1.10$ volts

Half Cells: Both parts of the cell are called half cells. No reaction can take place in any isolated half cell. Both the reactions take place simultaneously at both half cells. In this cell, electrons travel in the external circuit while ions travel through the porous partition in the internal circuit.

Production of Daniel Cell: Daniel cell produces a voltage of around 1.10 volts under standard conditions.

Q.21: How the reactivity of metals can be found out by using voltage data obtained from galvanic cell?

Ans. Order of Reactivity of Metals Using Voltage Data:

The reactivity of metals can be found out by using voltage data obtained from galvanic cells.

According to the available data the reactivity of metals depends upon how readily it is oxidized or lose electrons to form positive ions.

Establishment of a Galvanic Cell: If a galvanic cell is established between two metals, the more reactive metal will lose electrons to form positive ions.

Voltage Generated by a Galvanic cell: The voltage generated by a galvanic cell measures the comparative tendencies of the metals to lose electrons.

Higher voltage: A higher voltage indicates that the difference in the tendency of one metal (more reactive metal) to lose electrons is for more than the other.

Magnesium is more reactive than Iron: If a cell is established between magnesium and copper it produces a higher voltage than a cell established between iron and copper. This will mean that magnesium is more reactive than iron because it has a tendency to lose

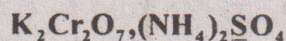
electrons more easily.

Electrochemical Series: By measuring the voltage produced when different metal pairs are used in a voltaic cell, it is possible to determine the reactivity series of metals which is called electrochemical series.

More Reactive Metals: Metals placed higher in this series are considered more reactive and easily lose electrons than the metals placed lower in the series.

16.1 Quick Check!

1. Assign oxidation number to the marked atom of the following compounds:



- (i) $K_2Cr_2O_7$ (Potassium dichromate)

Ans. $K = +1, O = -2, \text{ let } Cr = x$

- $2(+1) + 2(x) + 7(-2) = 0$
- $2 + 2x - 14 = 0$
- $2x = 12$
- $x = +6$
- Oxidation number of Cr is +6

- (ii) $(NH_4)_2SO_4$ (Ammonium Sulphate)

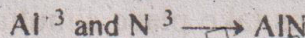
Ans. $S = x, O = -2$

- $x + 4(-2) = -2$
- $x - 8 = -2$
- $x = +6$
- Oxidation number of S is +6.

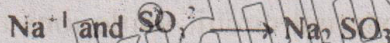
16.2 Quick Check!

1. Write the formulae of the following ionic compounds with the help of oxidation numbers of their constituent atoms. Aluminum nitride, Sodium sulfite and Sodium hydrogen carbonate.

Ans. Aluminum nitride:



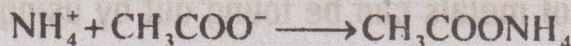
- Sodium Sulfite:



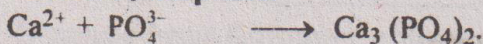
- Sodium Hydrogen Carbonate: $Na^{+1} \text{ and } HCO_3^{-} \longrightarrow NaHCO_3$

2. Write the formulae of following ionic compounds with the help of charges present on their ions. Ammonium acetate, Calcium phosphate and Potassium chromate.

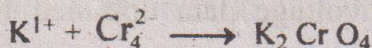
Ans. Ammonium acetate:



Calcium phosphate:



- Potassium chromate:



16.3 Quick Check!

1. Identify the oxidising agent and the reducing agent in the following reactions.

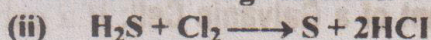


Ans. • Cu goes from +2 to +1 (Reduced).

Oxidising Agent: $CuSO_4$

- I goes from -1 to 0 (Oxidized)

Reducing Agent: KI



Ans. • Cl goes from 0 to -1 (Reduced)

Oxidising Agent: Cl_2

- S goes from -2 to 0 (Oxidized)

Reducing Agent: H_2S

16.4 Quick Check!

1. Why carbon electrodes are used in the electrolysis of concentrated solution of NaCl?

Ans. Carbon (graphite) electrodes are used because they are inert (do not react with the products like chlorine gas) and are excellent conductors of electricity.

16.5 Quick Check!

1. Which other electrolyte can be used in hydrogen-oxygen fuel cell?

Ans. Apart from alkaline solutions (like KOH), phosphoric acid or solid polymers (PEM) can be used as electrolysis in hydrogen-oxygen fuel cells.

16.6 Quick Check!

1. What conditions are required for rusting of iron.

Ans. Presence of oxygen (air) and moisture (water) are required for rusting.

2. If a sample of iron and a sample of zinc come into contact, which will corrode? Explain your answer.

Ans. **Iron Vs. Zinc:** Zinc will corrode. Zinc is more reactive (more electropositive) than iron. In a process called "Sacrificial protection", the more active metal (Zinc) oxidizes first to protect the less active metal (iron).

16.7 Quick Check!

1. What is the principle behind electroplating?

Ans. **Principle:** It is based on electrolysis. A metal coating is desposited on an object by reducing metal ions from an electrolyte solution using an electric current.

2. How electroplating is used to repair worn-out areas of metal objects?

Ans. **Repairing Metal Objects:** Electroplating can deposit a layer of metal (like chromium or Nickel) onto. Worn-out areas to restore the object's thickness, shape, and provide a protective, shiny finish.

16.8 Quick Check!

1. Compare an electrolytic and a galvanic cell.

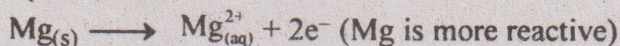
Ans.

Electrolytic Cell	Galvanic Cell
Converts electrical energy to chemical energy.	Converts chemical energy to electrical energy.
It is non-spontaneous reaction (requires a battery)	It is spontaneous reaction.

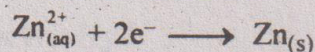
2. Write down the reactions which occur in a Zn – Mg galvanic cell.

Ans. Zn–Mg Galvanic cell reactions

• **Anode (Oxidation)**



• **Cathode (Reduction)**



• **Overall**



DO YOU KNOW?

- Spontaneous reactions occur naturally without any energy input, while non-spontaneous reactions require a continuous input of energy.

INTERESTING INFORMATION

- The terms anode and cathode always refer to the electrodes at which oxidation and reduction occur respectively. However the charges carried by electrodes in a battery and electrolytic cell are different.
- Oxidation numbers do not represent the real electrical charges present on atoms except in simple ions. However, the oxidation numbers are useful for keeping track of the redox reactions.
- A group of chemically unreactive and valuable metals are called noble metals. They include gold, silver and platinum.
- Electrolysis played a crucial role in the discovery and industrial production of elements like fluorine and aluminum.
- Chromium plating is usually done on iron or on steel. As chromium metal does not directly adhere to the iron properly, therefore, iron is first copper or nickel plated.
- The term oxidation state is also used interchangeably of oxidation number. In water, hydrogen has an oxidation number of +1 and it is said to be in the +1 oxidation state.

Examples

16.1 Assign oxidation number to sulphur in SO_4^{2-}

Sol. $\text{S} + 4\text{O} = -2$

$$\text{S} + 4(-2) = -2$$

$$\text{S} - 8 = -2$$

$$\text{S} = 8 - 2$$

$$\text{S} = +6$$

The oxidation number of S in SO_4^{2-} is +6.

16.2 Find out the oxidation number of manganese in KMnO_4 .

Sol. $\text{K} + \text{Mn} + 4\text{O} = 0$

$$(+1) + \text{Mn} + 4(-2) = 0$$

$$\text{Mn} - 7 = 0$$

$$\text{Mn} = +7$$

The oxidation number of Mn in KMnO_4 is +7.

16.3 Find out the oxidation number of Fe in Fe_2O_3 .

Sol. $2\text{Fe} + 3\text{O} = 0$

$$2\text{Fe} + 3(-2) = 0$$

$$2\text{Fe} - 6 = 0$$

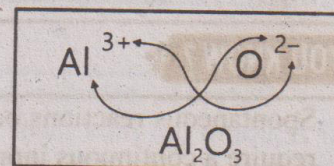
$$2\text{Fe} = +6$$

$$\text{Fe} = +3$$

Oxidation number of one atom of Fe is +3

16.4 Write the chemical formula of aluminum oxide with the help of oxidation numbers of the elements involved.

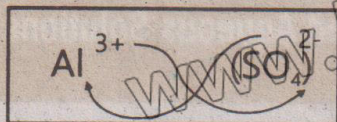
Sol. Aluminum is present in Group 13 so its principal oxidation number shall be +3. Oxygen being more electronegative atom normally shows -2 oxidation state in normal oxides. The formula of aluminum oxide will



then be Al_2O_3 because the some of oxidation numbers of both elements will be equal to zero with this formula.

16.5 Write down the formula for aluminum sulphate.

Sol.



The formula for aluminum sulphate is $\text{Al}_2(\text{SO}_4)_3$.

Objective Type Questions

Additional MCQ's

Multiple Choice Questions prepared in the light of new Examination Techniques
(Knowledge, Understanding, Application, Analytical & Conceptual)

16.1+16.2
+16.3+16.4
+16.5

Oxidation Number of Atoms in Compounds and Ions + Writing Chemical Formulae of Ionic Compounds by applying Oxidation Numbers + Writing Chemical formulae of Ionic compounds using ionic charges + Oxidation and Reduction + Oxidation - Reduction Reactions

- What is the oxidation number of hydrogen in a metal hydride like NaH ?
(a) +1 (b) -0 (c) -1 (d) +2
- The chemical process in which an atom or ion gains electrons is called
(a) Oxidation (b) Reduction (c) Electrolysis (d) Corrosion
- In a neutral compound, the sum of the oxidation numbers of all atoms is always:
(a) 0 (b) +1 (c) -1 (d) +2
- Which of the following is the correct formula for aluminum sulphate?
(a) AlSO_4 (b) Al_2SO_4 (c) $\text{Al}_3(\text{SO}_4)_2$ (d) $\text{Al}_2(\text{SO}_4)_3$
- What is the oxidation number of sulfur (s) in H_2SO_4 ?
(a) +4 (b) +6 (c) -2 (d) +2
- The movement of which of the following allows on aqueous solution to conduct electricity?
(a) Free electrons (b) Protons (c) Mobile ions (d) Newtrons
- In the reaction $\text{CuO} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{O}$, which substance is being oxidized?
(a) CuO (b) H_2 (c) Cu (d) H_2O
- Which of there is a common by-product of a hydrogen-oxygen fuel cell?
(a) Carbon dioxide (b) Sulfur dioxide (c) Water (d) Methane

**16.6 + 16.7
+ 16.8 + 16.9**

Electrolysis + Electrolysis of a Concentrated Aqueous Solution of Sodium Chloride (Brine) using inert electrodes (Graphite / Pt) + Electrolysis of Dilute and Concentrated Aqueous solutions of Metal Halides + Electrolysis of Concentrated Aqueous Solutions of Metal Halides

9. What is the process of decomposing a chemical compound into its constituents by passing an electric current through it?
(a) Neutralization (b) Evaporation (c) Ionization (d) Electrolysis
10. In an electrolytic cell, the electrode connected to the positive terminal of the battery is the:
(a) Electrolyte (b) Anode (c) Cathode (d) Anolyte
11. During electrolysis, reduction (gain of electrons) always occurs at the:
(a) Cathode (b) Anode (c) Battery (d) External circuit
12. In the electrolysis of concentrated sodium chloride (Brine), which gas is released at the anode?
(a) Nitrogen (N_2) (b) Oxygen (O_2) (c) Hydrogen (H_2) (d) Chlorine (Cl_2)
13. According to the electrochemical series, which of these metals is the strongest reducing agent?
(a) Gold (Au) (b) Lithium (Li) (c) Silver (Ag) (d) Nickel (Ni)
14. When an aqueous solution of copper (II) sulfate is electrolyzed with inert electrodes, what is deposited at the cathode?
(a) Sulfur dioxide (b) Copper metal (c) Hydrogen gas (d) Oxygen gas

**16.10 + 16.11
+ 16.12 + 16.13**

Electrolysis of Molten Lead (II) chloride + Electrolysis of Dilute solution of Sulphuric acid using inert electrodes (carbon or Pt) + Electrolysis of Dilute Aqueous Solution of Copper Sulphate using inert electrodes or Copper electrodes + Fuel Cells

15. At what temperature does lead (II) chloride melt to become an electrolyte?
(a) $100^\circ C$ (b) $501^\circ C$ (c) $801^\circ C$ (d) $373^\circ C$
16. During the electrolysis of molten $PbCl_2$, what is formed at the cathode?
(a) Chlorine gas (b) Oxygen gas (c) Molten lead (d) Hydrogen gas
17. In the electrolysis of dilute sulphuric acid, which ion is discharged at the anode?
(a) SO_4^{2-} (b) OH^- (c) H^+ (d) Cl^-
18. When aqueous $CuSO_4$ is electrolyzed using inert (carbon) electrodes, what gas is produced at the anode?
(a) Oxygen (b) Hydrogen (c) Carbon dioxide (d) Sulphur dioxide
19. A fuel cell converts chemical energy directly into:
(a) Heat energy (b) Mechanical energy
(c) Sound energy (d) Electrical energy

20. What is the by-product of a hydrogen-oxygen fuel cell?
 (a) Carbon dioxide (b) Water (c) Methane (d) Nitrogen oxide
21. Which catalyst is commonly used in the porous carbon electrodes of a fuel cell?
 (a) Copper (b) Zinc (c) Platinum (d) Iron

16.14+16.15
 +16.16+16.17

Corrosion + Electroplating + Galvanic cells + Order of
 Reactivity of Metals using Voltage Data

22. Which two substances are essentially required for the rusting of iron?
 (a) Nitrogen and water (b) Oxygen and water
 (c) Carbon dioxide and water (d) Hydrogen and water
23. Rusting is considered an electrochemical process where iron loses electrons to form:
 (a) Fe (III) ions (b) Fe (II) oxides (c) Oxide ions (d) Hydroxide ions
24. Galvanization involves coating iron with a thin layer of which metal?
 (a) Copper (b) Zinc (c) Gold (d) Silver
25. In sacrificial protection, which metal is often used to protect buried fuel tanks?
 (a) Magnesium (b) Tin (c) Lead (d) Copper
26. During electroplating, the object to be plated is always made the:
 (a) Cathode (b) Anode (c) Electrolyte (d) Salt bridge
27. In a Daniel cell, the zinc electrode acts as the:
 (a) Electrolyte (b) Porous partition (c) Anode (d) Cathode
28. In the electroplating of copper, the electrolyte used is:
 (a) Copper sulphate (b) Silver nitrate
 (c) Sodium chloride (d) Potassium cyanide
29. Under standard conditions, Daniel cell produced a voltage of around:
 (a) 1.10 volts (b) 1.20 volts (c) 1.30 volts (d) 1.40 volts
30. A Galvanic cell (like the Daniel cell) converts the chemical energy into:
 (a) Sound energy (b) Electrical energy (c) Light energy (d) Heat energy

Answers

1. (c) 2. (b) 3. (a) 4. (d) 5. (b) 6. (c) 7. (b) 8. (c)
 9. (d) 10. (b) 11. (a) 12. (d) 13. (b) 14. (b) 15. (b) 16. (c)
 17. (b) 18. (a) 19. (d) 20. (b) 21. (c) 22. (b) 23. (a) 24. (b)
 25. (a) 26. (a) 27. (c) 28. (a) 29. (a) 30. (b)

Additional Short Questions

Short Answer Questions prepared in the light of new Examination Techniques
(Knowledge, Understanding, Application, Analytical & Conceptual)

16.1 + 16.2
+ 16.3 + 16.4
+ 16.5

Oxidation Number of Atoms in Compounds and Ions + Writing Chemical Formulae of Ionic Compounds by applying Oxidation Numbers + Writing Chemical formulae of Ionic compounds using ionic charges + Oxidation and Reduction + Oxidation - Reduction Reactions

1. Define oxidation number.

Ans. Oxidation number: The oxidation number is an apparent charge on an atom in a compound or an ion. It may have a positive or a negative sign.

2. How does oxidation number help us?

Ans. Oxidation number helps us:

- To identify whether the reaction involves oxidation and reduction or not.
- To identify oxidizing and reducing agents in a given oxidation-reduction reaction.
- To write chemical formulas and names of the chemical compounds.
- To keep track of the new distribution of electrons during an oxidation-reduction reaction.

3. What is the oxidation numbers of any element in its free (uncombined) state?

Ans. The oxidation number of an atom in a free element is always zero (e.g; Na, O₂ or Cl₂ all have an oxidation of 0)

4. What is the sum of oxidation numbers in a neutral compound?

Ans. The algebraic sum of oxidation numbers of all atoms in a neutral compound is always zero.

5. How is the oxidation number assigned to a monoatomic ion?

Ans. The oxidation number is equal to the actual charge present on the ion (e.g; Ca²⁺ is +2 and Cl⁻ is -1)

6. Define oxidation in terms of electrons.

Ans. Oxidation is a process in which electron or electrons are lost.

7. Define reduction in terms of electrons.

Ans. Reduction is a process in which electron or electrons are gained.

8. What is a redox reaction?

Ans. Redox Reaction: A chemical reaction which involves simultaneous oxidation and reduction is called an oxidation-reduction or redox reaction.

The reaction may involve the complete transference of electron or electrons to form ionic bonds or a partial transfer or shift of electron to form covalent bonds.

9. What is an oxidizing agent?

Ans. An oxidizing agent is a substance that gains electrons and gets reduced itself. It causes the oxidation of another substance.

10. What is a reducing agent?

Ans. A reducing agent is a substance that loses electrons and gets oxidized itself. It causes the reduction of another substance.

11. Why does the purple colour of potassium permanganate (KMnO_4) disappear in a reaction with oxalic acid?

Ans. The purple colour disappears because the Mn^{+7} (permanganate ion) is reduced to Mn^{+2} (manganese ion), which is colourless in dilute solution.

12. What is a polyatomic ion?

Ans. Polyatomic ion: A group of atoms that are bonded together and carry an overall not charge (e.g., carbonate CO_3^{2-} or Ammonium NH_4^+).

16.6+16.7 Electrolysis + Electrolysis of a Concentrated Aqueous Solution of Sodium Chloride (Brine) using inert electrodes (Graphite / Pt) + Electrolysis of Dilute and Concentrated Aqueous solutions of Metal Halides + Electrolysis of Concentrated Aqueous Solutions of Metal Halides

13. Define electrolysis.

Ans. Electrolysis: The process in which an electricity is passed through fused ionic compounds or their aqueous solution, they are decomposed into their constituents, is called electrolysis.

14. What is an electrolytic cell?

Ans. Electrolytic Cell: An electrolytic cell is an apparatus consisting of a vessel (electrolytic both), an electrolyte, and two electrodes connected to an external voltage source (battery) to carry out electrolysis.

15. In What the process of electrolysis can be carried out?

Ans. The process of electrolysis can be carried out in an electrolytic both called an electrolytic cell with the help of two electrodes through which electric current enters and leaves the cell.

16. Define the anode and cathode in an electrolytic cell.

Ans. Anode: It is a positive electrode through which electrons enter the external circuit.

Cathode: It is a negative electrode through which electrons leave the external circuit.

17. Why must can electrolyte must be in fused (molten) state or in solution form?

Ans. In solid form, ions are held tightly in a lattice and cannot move. In a fused or aqueous state, the ions are free to move toward the electrodes to conduct electricity.

18. What is Brine?

Ans. Brine: Brine is a concentrated aqueous solution of sodium chloride (NaCl).

19. Which gases are produced during the electrolysis of Brine?

Ans. During the electrolysis of Brine, following gases are produced:

At the cathode: Hydrogen gas (H_2)

At the Anode: Chlorine gas (Cl_2)

- 20. What determines which ion will be discharged at an electrode when multiple ions are present?**

Ans. The discharged depends on the electrochemical series. Elements with a higher tendency to lose or gain electrons (based on their position in the series) will be preferentially discharged.

- 21. If a metal is below hydrogen in the electrochemical series, what is discharged at the cathode?**

Ans. If a metal is below hydrogen in the electrochemical series, then the metal ion is discharged.

- 22. If a metal is above hydrogen in the electrochemical series, what is produced at the cathode?**

Ans. If a metal is above hydrogen in the electrochemical series, then hydrogen ion is discharged.

- 23. Why is the solution around the cathode becomes alkaline during the electrolysis of brine?**

Ans. The solution around the cathode becomes alkaline during the electrolysis of brine due to the formation of hydroxide ions (OH^-).

- 24. Why is platinum sometimes used as an electrode?**

Ans. Platinum is an inert electrode, meaning it conducts electricity but does not chemically react with the electrolyte or the products (like chlorine gas) during the process.

- 25. Mention two industrial applications of electrolysis.**

Ans. (i) Production of elements like Aluminum, sodium and fluorine.

(ii) Manufacturing of chemicals like sodium hydroxide ($NaOH$).

**16.10+6.11
+16.12+16.13**

Electrolysis of Molten Lead (II) chloride + Electrolysis of Dilute solution of Sulphuric acid using inert electrodes (carbon or Pt) + Electrolysis of Dilute Aqueous Solution of Copper Sulphate using inert electrodes or Copper electrodes + Fuel Cells

- 26. Why must lead (II) chloride be heated to $501^\circ C$ before electrolysis can occur?**

Ans. Lead chloride (II) is a solid at room temperature. Heating it to $501^\circ C$ melts it, making it an electrolyte where ions are free to move and conduct electricity.

- 27. What happens to the concentration of sulphuric acid as electrolysis proceeds?**

Ans. As the electrolysis proceeds, the concentration of sulphuric acid increases because the water present is being decomposed.

28. What is deposited at the cathode during the electrolysis of dilute aqueous solution of copper sulphate using inert electrodes?

Ans. Copper metal is deposited at the cathode (inert) during the electrolysis of aqueous solution of copper sulphate using inert electrodes.

29. Which gas is produced during the electrolysis of dilute aqueous solution of copper sulphate using inert electrodes?

Ans. Oxygen gas is produced at anode (inert) during the electrolysis of dilute aqueous solution of copper sulphate using inert electrodes.

30. Why the blue colour of the solution fades away during the electrolysis of dilute aqueous solution of copper sulphate using inert electrodes?

Ans. During the electrolysis of dilute aqueous solution of copper sulphate using inert electrodes, since the concentration of copper ions decreases slowly so the blue colour of the solution fades away with the passage of time.

31. How does the process change when copper electrodes are used instead of inert electrodes in CuSO_4 solution?

Ans. If copper electrodes are used instead of inert electrodes, the copper metal atoms from the copper anode start moving into the solution in the form of copper ions and the electrode mass starts decreasing.

32. Define a fuel cell.

Ans. Fuel Cell: It is a device that provides a method in which chemical energy is converted to electrical energy. It uses gaseous fuels like hydrogen and oxygen to produce electricity.

33. What are the electrodes in a hydrogen-oxygen fuel cell usually made of?

Ans. The electrodes are made of porous compressed carbon saturated with platinum which acts as a catalyst.

34. Why is a fuel cell considered highly efficient compared to traditional engines?

Ans. A fuel cell is very efficient because it converts 75% of the bond energy of fuels to electricity, whereas traditional heat engines lose much more energy as heat.

16.14+16.15
+16.16+16.17

Corrosion + Electroplating + Galvanic cells + Order of
Reactivity of Metals using Voltage Data

35. Define corrosion and give a common example.

Ans. Corrosion: Corrosion is a natural process due to which a metal suffers gradual decay when it interacts chemically with its environment.

Example: A common example is rusting of iron.

36. Why does iron continue to corrode until it is completely destroyed?

Ans. Once iron begins to rust, the process will continue as rust is porous and allows both air and moisture to come into contact with fresh metal underneath. Hence, all the

metal will corrode given enough time.

37. What is Galvanization?

Ans. Galvanization: Galvanization is a common method to protect iron or steel from rusting. It involves the dipping of an iron sheet in a bath of molten zinc. A layer of zinc metal gets deposited on the iron sheet and prevents rusting.

Example: Tin cans are actually steel cans which have been coated with a thin layer of tin.

38. Define electroplating.

Ans. Electroplating: It is a process in which a metal is deposited on another metal electrolytically.

39. What properties should a chosen electrolyte have?

Ans. The electrolyte chosen should have the following properties:

- Very soluble in water.
- Good conductor.
- Not easily oxidized or reduced or hydrolyzed.
- Inexpensive.

40. What conditions should be fulfilled to get good deposits on the electroplates during the electroplating?

Ans. The following conditions should be fulfilled to get good deposits.

- High current density
- Low temperature
- High concentration of the metal in its electrolyte.

41. What does Daniel cell use to produce electric current?

Ans. Daniel cell uses oxidation-reduction reactions to produce electric current.

42. In the electroplating process, which electrode is the object to be plated?

Ans. The article or object to be plated is always made the cathode (the negative electrode), while the metal to be deposit is made the anode.

43. What is a Galvanic (Voltaic) cell?

Ans. Galvanic Cell: A galvanic cell is an electrochemical cell that converts chemical energy into electrical energy through spontaneous oxidation-reduction (redox) reactions.

44. What is the function of the porous partition in a Daniel cell?

Ans. The porous partition in a Daniel cell keeps the two electrolyte solutions (Zinc Sulphate and Copper Sulphate) separate but allows ions to pass through to complete the electrical circuit.

45. How can voltage data be used to determine the reactivity of metals?

Ans. A higher voltage generated in a galvanic cell indicates a greater difference in the reactivity of the two metals used. Metals that lose electrons more easily (more reactive) produce higher voltages when paired with a standard electrode.

A. Multiple Choice Questions

- ☆ Choose and tick the correct answer from the given choices.
- The oxidation number of Mn in K_2MnO_4 :
 (a) +7 (b) +5 (c) +6 (d) -6
 - Which element in the following reaction is being reduced?
 $4HBr_{(aq)} + MnO_{2(s)} \longrightarrow MnBr_{2(aq)} + Br_{2(g)} + 2H_2O_{(l)}$
 (a) Br (b) Mn (c) H (d) Both Mn and H
 - During electrolysis of dilute solution of H_2SO_4 , using Pt electrodes, which process takes place at the cathode?
 (a) Reduction (b) Neither oxidation nor reduction
 (c) Oxidation (d) First oxidation and then reduction
 - Which product you will get at the anode when an aqueous solution of $CuSO_4$ is electrolysed using copper electrodes?
 (a) Cu (b) O_2 (c) H_2
 (d) No product, instead copper metal will go into the solution as Cu^{2+} .
 - Which of the following statements is NOT correct about Zn-Cu electrochemical cell?
 (a) Anode is negatively charged (b) Reduction occurs at anode
 (c) Cathode is positively charged (d) Reduction occurs at cathode
 - Which product will be obtained at the anode when a concentrated solution of NaCl will be electrolysed?
 (a) Cl_2 (b) O_2
 (c) More O_2 and less Cl_2 (d) More Cl_2 and less O_2
 - When a dilute aqueous solution of $ZnCl_2$ is electrolysed which product you will get at the cathode?
 (a) H_2 (b) Zn (c) O_2 (d) Zn and O_2
 - What is produced at the cathode in hydrogen-oxygen fuel cell?
 (a) H^+ (b) O^{2-} (c) H_2O (d) OH^{1-}
 - Which of the following metals is highest in the electrochemical series?
 (a) Tin (b) Iron (c) Magnesium (d) Zinc

Answers

1. (c) 2. (b) 3. (a) 4. (d) 5. (b) 6. (a) 7. (a)
 8. (d) 9. (c)

B. Short Answer Questions

16.1 What happens when electricity is passed through an aqueous solution of NaCl?

Ans. When electricity passes through aqueous solution of NaCl (electrolysis), the salt dissociates into Na^+ and Cl^- ions.

At the Anode: Cl^- ions lose electrons to form chlorine gas (Cl_2).

At the cathode: H^+ ions (from water) gain electrons to form hydrogen gas (H_2).

In solution: Na^+ and OH^- ions remain, forming sodium hydroxide (NaOH).

16.2 What are the main objectives of electroplating?

Ans. For answer see Q18 (Long Subjective Questions). First part

16.3 Mention one difference between an electrolytic cell and a galvanic cell.

Ans.

Electrolytic Cell	Galvanic Cell
Converts electrical energy to chemical energy.	Converts chemical energy to electrical energy.
It is non-spontaneous reaction (requires a battery)	It is spontaneous reaction

16.4 Through which electrode, electrons leave the electrolytic cell?

Ans. Electrons leave the electrolytic cell through the anode.

Reason: Electrons flow from the external power source (battery) into the cathode, and they exit the electrolyte through the anode to return to the positive terminal of the battery.

16.5 How does an electrolyte conduct electricity?

Ans. An electrolyte conducts electricity through the movements of ions.

- When a salt or acid is dissolved in water or melted, it dissociates into positive ions (cations) and negative ions (anions).
- When an electric current is applied, the cations move toward the negative electrode (cathode) and the anions move toward the positive electrode (anode), allowing the current to flow through the liquid.

16.6 What will happen if a strip of Cu metal is dipped in FeSO_4 solution?

Ans. If a strip of Cu metal is dipped in FeSO_4 solution, no reaction will occur.

Reason: In the reactivity series, copper (Cu) is less reactive than iron (Fe). Because copper is lower in the series, it cannot displace iron from its sulfate solution.

16.7 Why fuel cells are regarded as environment friendly?

Ans. Fuel cells are regarded as environmental friendly because of two main reasons:

- **Zero Harmful Emissions:** Unlike internal combustion engines, the only by-product of a hydrogen-oxygen fuel cell is pure water and heat.
- **Higher efficiency (less fuel waste):**
Fuel cells convert chemical energy directly into electrical energy without combustion, so they are more efficient than conventional engines. This results in less fuel consumption and reduced environmental impact.

C. Constructed Response Questions

16.1 Calculate the oxidation number of chlorine atom in $\text{Ca}(\text{ClO}_3)_2$.

Ans. To find the oxidation number of chlorine, we use the known oxidation number for calcium and oxygen.

- **Calcium (Ca):** +2 (Group 2 element)
- **Oxygen (O):** -2 (usually)
- **Chlorine (Cl):** x

The sum of oxidation numbers in a neutral compound is 0.

$$1(+2) + 2[x + 3(-2)] = 0$$

$$2 + 2[x - 6] = 0$$

$$2 + 2x - 12 = 0$$

$$2x - 10 = 0$$

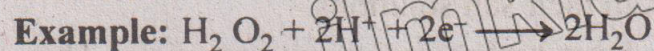
$$2x = +10$$

$$x = +5$$

The oxidation number of chlorine in $\text{Ca}(\text{ClO}_3)_2$ is +5.

16.2 How can H_2O_2 behave as an oxidizing agent?

Ans. Hydrogen peroxide (H_2O_2) acts as an oxidizing agent because the oxygen atom in it is an intermediate oxidation state of -1. It can easily accept electrons to reach the more stable -2 state (as found in water).



16.3 Why different products are formed at the electrodes when a dilute and concentrated aqueous solutions of NaCl are electrolysed?

Ans. This happens due to the selective discharge of ions, which depends on their concentration.

In Dilute NaCl: At the anode, OH^- ions are in higher concentration and are discharged to produce oxygen gas.

In Concentrated NaCl (Brine): At the anode, the high concentration of Cl^- ions resulting in the production of chlorine gas.

16.4 Why corrosion becomes fast during rainy season?

Ans. Corrosion (rusting) is an electrochemical process that requires water and oxygen.

During the rainy season:

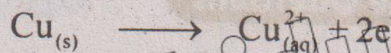
- The high moisture/humidity provides the necessary electrolyte for the movement of ions.
- Rainwater often contains dissolved gases (like CO_2) that make it slightly acidic, increasing its conductivity and accelerating the redox reaction.

16.5 If a galvanic cell is established between Cu and Ag, what reactions do you expect at the electrodes?

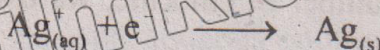
Ans. In the electrochemical series, copper (Cu) is more reactive than silver (Ag).

Therefore, Cu will be the anode and Ag will be the cathode.

- At Anode (Oxidation)



- At Cathode (Reduction)



16.6 What is the function of a porous partition in a galvanic cell?

Ans. The porous partition (or salt bridge) serves two main purposes.

- **Completes the Circuit:** It allows the flow of ions between the two half-cells.
- **Maintains Electrical Neutrality:** It prevents the accumulation of charges in the half-cells by allowing ions to migrate, ensuring the cell continues to function.

16.7 How do fuel cells produce electricity?

Ans. For answer see Q14 (Long Subjective Questions)

16.8 How can the chemical reactivity of metals be compared by using electrochemical series?

Ans. The electrochemical series ranks metals based on their standard reduction potentials:

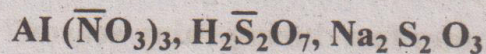
- **Higher in the series:** Metals have lower reduction potentials (or higher oxidation potentials). They lose electrons easily, are more reactive, and can displace metals below them from their salt solutions.
- **Lower in the series:** Metals have higher reduction potentials. They are less reactive and act as better oxidizing agents.

D. Descriptive Questions

16.1 Explain with examples, how the formula of ionic compound can be written using oxidation numbers.

Ans. For Answer see Q2 (Long Subjective Questions)

16.2 Find the oxidation numbers of the indicated elements in the following compounds.



Ans. To find these, we use the rule that the sum of oxidation number in a neutral compound is 0. Common oxidation numbers used here:

$$\text{H} = +1, \quad \text{O} = -2, \quad \text{Na} = +1, \quad \text{Al} = +3$$

Al($\overline{\text{NO}}_3$)₃: Since the nitrate ion (NO_3^-) has a charge of -1, we can calculate N within that ion:

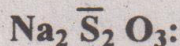
- Let $\text{N} = x$
- $x + 3(-2) = -1$
- $x - 6 = -1$
- $x = +5$

Oxidation number of N is +5.



- $2(\text{H}) + 2(\text{S}) + 7(\text{O}) = 0$
- $2(+1) + 2(x) + 7(-2) = 0$
- $2 + 2x - 14 = 0$
- $2x - 12 = 0$
- $2x = 12$
- $x = +6$

Oxidation number of S is +6.



- $2(\text{Na}) + 2(\text{S}) + 3(\text{O}) = 0$
- $2(+1) + 2(x) + 3(-2) = 0$
- $2 + 2x - 6 = 0$
- $2x - 4 = 0$
- $2x = 4$
- $x = +2$

16.3 Write down the formulae of the following compounds using ionic charges.

Sodium hypochlorite, calcium oxalate, ammonium phosphate and calcium nitrite

Ans. To write the formulae of the following compounds using ionic charges, we balance the positive and negative charges so the final compound is neutral.

Compound Name	Ions Involved	Chemical Formula
Sodium hypochlorite	Na^+ and ClO^-	Na ClO
Calcium oxalate	Ca^{2+} and $\text{C}_2\text{O}_4^{2-}$	$\text{Ca C}_2\text{O}_4$
Ammonium Phosphate	NH_4^+ and PO_4^{3-}	$(\text{NH}_4)_3\text{PO}_4$
Calcium nitrite	Ca^{2+} and NO_2^-	$\text{Ca}(\text{NO}_2)_2$

16.4 Describe the working of a fuel cell.

Ans. For Answer see Q 14

16.5 Discuss the electrolysis of concentrated solutions of metal halides.

Ans. For Answer see Q 10

16.6 Explain electroplating giving examples.

Ans. For Answer see Q 18

16.7 What is corrosion? Explain the reactions involved in corrosion of iron.

Ans. For Answer see Q 16

