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CHAPTER 16

ELECTROCHEMISTRY

Class 10 — Chemistry Notes

(Extracted & Compiled — Branded for taleemzone.com)

Contents

DESCRIPTIVE QUESTIONS

OXIDATION NUMBER OF ATOMS IN COMPOUNDS AND IONS

Q.1 Define oxidation number. Why is it important to assign an oxidation number to an atom? *[10216001]*

Oxidation Number of Atoms in Compounds and Ions

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.

Importance of Oxidation Number

It is important to assign an oxidation number to an atom in a compound because of following reasons.

- **(i)** The assignment of oxidation number to atoms in compounds and ions helps us to identify whether the reaction involves oxidation and reduction or not.
- **(ii)** It also helps to identify oxidizing and reducing agents in a given oxidation-reduction reaction.
- **(iii)** Oxidation numbers are also used to write chemical formulas and names of the chemical compounds.
- **(iv)** This concept also helps to keep track of the new distribution of electrons during an oxidation-reduction reaction.

Q.2 Write the rules to find out the oxidation number of atoms in compounds. *[10216002]*

Rules for Finding Oxidation Numbers

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

- **(i) Zero oxidation states:** 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.
- **(ii) Monoatomic ions:** 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.
- **(iii) Elements of groups 1, 2 and 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.
- **(iv) Binary compounds:** In binary compounds the more electronegative atom is always assigned a negative oxidation number. 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.

Point to note — S.Q. What is the role of oxidation number in redox reactions? [10216003]
Ans. 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.

- **(v) Elements of group 17:** 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.
- **(vi) Oxidation state of hydrogen:** 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.
- **(vii) Oxidation number of oxygen:** The element oxygen is assigned an oxidation number -2 in most of its compounds.
- **(viii) Sum of oxidation numbers in neutral compounds:** The sum of oxidation numbers of all the atoms present in a neutral compound is always zero. In polyatomic ions: In a polyatomic ion the sum of oxidation numbers is equal to the charge present on that ion.

16.1 Example

Assign oxidation number to sulphur in SO_4^{2-}. *[10216004]*

Solution

$$S + 4\,O = -2$$
$$S + 4(-2) = -2$$
$$S - 8 = -2$$
$$S = 8 - 2$$
$$\mathbf{S = +6}$$

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

Point to note — S.Q. How oxidation number is related to oxidation state? [10216005]
Ans. 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.

16.2 Example

Find out the oxidation number of manganese in $KMnO_4$. *[10216006]*

Solution

$$K + Mn + 4O = 0$$
$$(+1) + Mn + 4(-2) = 0$$
$$Mn - 7 = 0$$
$$\mathbf{Mn = +7}$$

The Oxidation number of Mn in $KMnO_4$ is +7.

16.3 Example

Find out the oxidation number of Fe in Fe_2O_3. *[10216007]*

Solution

$$2Fe + 3O = 0$$
$$2Fe + 3(-2) = 0$$
$$2Fe - 6 = 0$$
$$2Fe = +6$$
$$\mathbf{Fe = +3}$$

Oxidation number of one atom of Fe is +3.

Quick Check 16.1
1. Assign oxidation number to the marked atom of the following compounds:
$K_2\underline{Cr}_2O_7$, $(NH_4)_2\underline{S}O_4$ *[10216008]*
Ans.

i. $K_2Cr_2O_7$	ii. $(NH_4)_2SO_4$
O.N of K = +1 O.N of O = -2 O.N of Cr = ? = x 2(+1) + 2(x) + 7(-2) = 0 2 + 2x - 14 = 0	O.N of (NH_4) = +1 O.N of O = -2 O.N of S = ? = x 2(+1) + x + 4(-2) = 0 2 + x - 8 = 0

2x = 12, x = +6	x = +6
The oxidation number of Cr in $K_2Cr_2O_7$ is +6.	The oxidation number of S in $(NH_4)_2SO_4$ is +6.

WRITING CHEMICAL FORMULAS OF IONIC COMPOUNDS BY APPLYING OXIDATION NUMBERS

Q.3 Explain with examples how formula of an ionic compound can be written using oxidation numbers. *[10216009]*

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 element calcium is present in group II (group 2) 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 upto zero.

$$\mathbf{Ca^{2+}\quad Cl^{1-}}$$

$$\mathbf{CaCl_2}$$

Example of Aluminium Oxide (16.4 Example)

[10216010]

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

Solution

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 sum of oxidation numbers of both elements will be equal to zero with this formula.

$$\mathbf{Al^{3+}\quad O^{2-}}$$

$$\mathbf{Al_2O_3}$$

WRITING CHEMICAL FORMULAS OF IONIC COMPOUNDS USING IONIC CHARGES

Q.4 Explain with example how chemical formula of an ionic compound having polyatomic ion can be written using ionic charges. *[10216011]*

Writing Chemical Formulas of Ionic Compounds using Ionic Charges

$$\mathbf{Cation:\ K^{1+}\qquad Anion:\ MnO_4^{1-}}$$

$$\mathbf{KMnO_4}$$

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

The formula for ammonium carbonate will be written as follows.

$$NH_4^{1+} \quad CO_3^{2-}$$

$$(NH_4)_2CO_3$$

The formula for ammonium carbonate will, therefore, be $(NH_4)_2CO_3$.

16.5 Example

[10216012]

Write down the formula of aluminium sulphate.

$$Al^{3+} \quad (SO_4)^{2-}$$

The formula for aluminum sulphate is $Al_2(SO_4)_3$.

Examples of several polyatomic ions are given in table 16.1

Table 16.1 — Common Polyatomic Ions					
CH3COO−	acetate	Cr2O72−	dichromate	NO2−	nitrite
NH4+	ammonium	HCO3−	bicarbonate	C2O42−	oxalate
CO32−	carbonate	HSO4−	bisulphate	ClO4−	perchlorate
ClO3−	chlorate	HSO3−	bisulphite	MnO4−	permanganate
ClO2−	chlorite	OH−	hydroxide	PO43−	phosphate
CrO42−	chromate	ClO−	hypochlorite	SO42−	sulphate
CN−	cyanide	NO3−	nitrate	SO32−	sulphite

Quick Check 16.2

1. Write the formulae of the following ionic compounds with the help of oxidation numbers of their constituent atoms. Aluminum nitride, Sodium sulphite and Sodium hydrogen carbonate. [10216013]

Ans. i. Aluminium nitride: $Al^{3+} + N^{3-} \rightarrow AlN$. ii. Sodium sulphite: $Na^+ + SO_3^{2-} \rightarrow Na_2SO_3$. iii. Sodium hydrogen carbonate: $Na^+ + HCO_3^- \rightarrow NaHCO_3$.

2. Write the formulae of following ionic compounds with the help of charges present on their ions. Ammonium acetate, Calcium phosphate and Potassium chromate. [10216014]

Ans. i. Ammonium acetate: $NH_4^+ + CH_3COO^- \rightarrow CH_3COONH_4$. ii. Calcium phosphate: $Ca^{2+} + PO_4^{3-} \rightarrow Ca_3(PO_4)_2$. iii. Potassium chromate: $K^+ + CrO_4^{2-} \rightarrow K_2CrO_4$.

OXIDATION AND REDUCTION

Q.5 Differentiate between oxidation and reduction by giving examples. *[10216015]*

Difference between Oxidation and Reduction

Oxidation and reduction are simultaneous but opposite processes. They always occur together. Oxidation and reduction changes are described in three principal ways.

Oxidation	Reduction
1. Oxidation is defined as a process in which oxygen is either added or hydrogen is removed. Examples (a) 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. $4Fe + 3O_2 \rightarrow 2Fe_2O_3$ $2Mg + O_2 \rightarrow 2MgO$	1. Reduction is defined as either addition of hydrogen or removal of oxygen. Examples (a) In the following reactions hydrogen is heated with metal oxide when oxygen comes out of the compound and combines with hydrogen to give a molecule of water. $CuO + H_2 \rightarrow$(Heat) $Cu + H_2O$ In the above examples, Cl_2 and CuO are being reduced and hydrogen (in both cases) is

Oxidation	Reduction
In the above reactions Fe and Mg are being oxidized while oxygen is being reduced. Oxygen is an oxidizing agent while iron and magnesium are reducing agents. (b) Hydrogen is removed from a compound by oxygen. Hydrogen is removed from CH_4 and HCl by oxygen or a compound containing oxygen.	being oxidized. Cl_2 and CuO are oxidizing agents while hydrogen in both the reactions will be reducing agent. (b) Hydrogen is added in chlorine to give hydrogen chloride gas. $H_2 + Cl_2 \rightarrow 2HCl$
$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ $4HCl + MnO_2 \rightarrow MnCl_2 + Cl_2 + 2H_2O$ 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 oxidizing agents.	
2. Loss of electron Oxidation is a process in which electron or electrons are lost. 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 a reducing agent. $Cu \rightarrow Cu^{1+} + e^-$ $Cu \rightarrow Cu^{2+} + 2e^-$	2. Gain of electrons Reduction is a process in which electron or electrons are gained. Example $Fe^{3+} + e^- \rightarrow Fe^{2+}$ $Mn^{7+} + 5e^- \rightarrow Mn^{2+}$ In both the above examples, ions of iron and manganese are being reduced. Hence, they are called oxidising agents.
3. Increase in oxidation state Oxidation is process in which oxidation number or oxidation state of an element increases. $Cu \rightarrow Cu^{1+} + e^-$ $Cu \rightarrow Cu^{2+} + 2e^-$ In the above two processes, the oxidation number of copper has increased from zero to +1 [Cu(I)] and +2 [Cu(II)] respectively.	3. Decrease in oxidation state Reduction is a process in which oxidation number of an element decreases. $Fe^{3+} + e^- \rightarrow Fe^{2+}$ $Mn^{7+} + 5e^- \rightarrow Mn^{2+}$ 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)].

OXIDATION – REDUCTION REACTIONS

Q.6 What are oxidation – reduction or redox reactions? Explain with examples. *[10216016]*

Oxidation – Reduction Reactions

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

(i) Reaction of $KMnO_4$ with $H_2C_2O_4$

Potassium permanganate acts as a very powerful oxidizing agent in acidic medium. It oxidizes oxalic acid into carbon dioxide.

Fig 16.1: Decolourization of KMnO4

Fig 16.1: Decolourization of KMnO4

Reaction

$$2KMnO_4 + 3H_2SO_4 \rightarrow K_2SO_4 + 2MnSO_4 + 3H_2O + 5[O]$$

$$[O] + H_2C_2O_4 \rightarrow 2CO_2 + H_2O \quad \times 5$$

$$2KMnO_4 + 3H_2SO_4 + 5H_2C_2O_4 \rightarrow K_2SO_4 + 2MnSO_4 + 10CO_2 + 8H_2O$$

Oxidising and Reducing Agent

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

$$Mn^{+7} + 5e^- \rightarrow Mn^{+2} \quad \text{(Reduction)}$$

$$C^{+3} - e^- \rightarrow C^{+4} \quad \text{(Oxidation)}$$

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

(ii) Reaction of $KMnO_4$ with KI

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.

Reaction

$$2KMnO_4 + 8H_2SO_4 + 10KI \rightarrow 6K_2SO_4 + 2MnSO_4 + 5I_2 + 8H_2O$$

$$Mn^{+7} + 5e^- \rightarrow Mn^{+2} \quad \text{(Reduction)}$$

$$2I^{-1} - 2e^- \rightarrow I_2^0 \quad \text{(Oxidation)}$$

Quick Check 16.3 — Identify the oxidizing agent and the reducing agent in the following reactions. [10216017]
(i) 2CuSO4 + 4KI → 2CuI + I2 + 2K2SO4
$Cu^{2+} \rightarrow Cu^{+}$ (Reduction) $I^- \rightarrow I_2^0$ (Oxidation)
In this reaction, copper sulphate (CuSO4) is oxidizing agent and potassium iodide (KI) is reducing agent.

(ii) H2S + Cl2 → S + 2HCl
$S^{2-} \rightarrow S^0$ (oxidation) $Cl2^0 \rightarrow Cl^-$ (reduction)
In this reaction, Cl2 (chlorine) acts as an oxidizing agent and hydrogen sulphide (H2S) acts as a reducing agent.

ELECTROLYSIS

Q.7 Define and explain general process of electrolysis. *[10216018]*

Electrolysis

When electricity is passed through fused ionic compounds or their aqueous solutions, they are decomposed into their constituents. Such compounds are called electrolytes and the process is called electrolysis.

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

- **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.

Fig 16.2: Electrolysis in an Electrolytic Cell

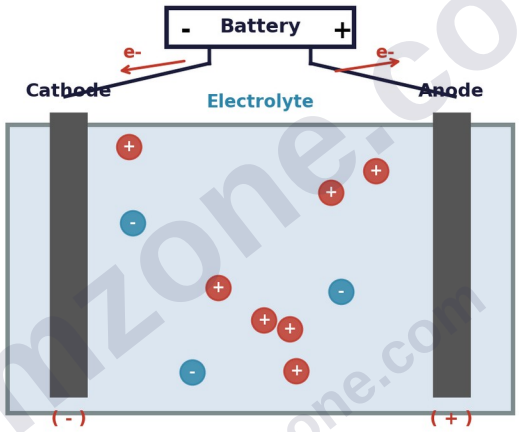

Fig 16.2: Electrolysis in an Electrolytic Cell

General Process

- 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.

THINGS TO KNOW! — S.Q What is difference between cathode and anode in battery and electrolytic cell? [10216019]
Ans. Difference between cathode and anode of battery and electrolytic cell (see table below).

	Electrode	Function	Charge
Battery	Anode	Oxidation	Negative

Battery	Cathode	Reduction	Positive
Electrolytic Cell	Anode	Oxidation	Positive
Electrolytic Cell	Cathode	Reduction	Negative

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.

ELECTROLYSIS OF A CONCENTRATED AQUEOUS SOLUTION OF SODIUM CHLORIDE (BRINE) USING INERT ELECTRODES (GRAPHITE/PT)

Q.8 Explain electrolysis of a concentrated aqueous solution of sodium chloride (Brine) using inert electrodes (Graphite/Pt). *[10216020]*

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 by-product.

- **Cathode:** The cathode used in this cell is made of platinum.
- **Anode:** The anode must be made of graphite to avoid the reaction with chlorine gas.

The following ions are present in the solution.

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

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

Fig 16.3: Electrolysis of Brine (Concentrated NaCl)

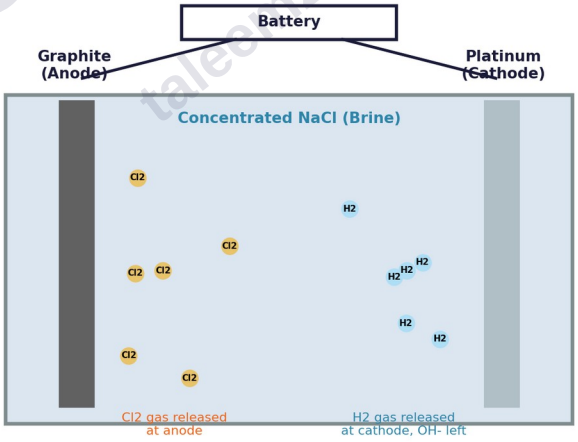

Fig 16.3: Electrolysis of Brine

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.

Reaction at Cathode

$$2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$$

or

$$2H_2O + 2e^- \rightarrow H_{2(g)} + 2OH^-$$

Did You Know? — Q. How Spontaneous and non-spontaneous reactions proceeds? [10216021]
Ans. Spontaneous reactions occur naturally without any energy input, while non-spontaneous reactions require a continuous input of energy.

Alkaline Solution around 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^-).

Reaction at Anode

$$2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$$

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:

Overall Reaction

$$2NaCl_{(aq)} + 2H_2O_{(l)} \rightarrow H_{2(g)} + Cl_{2(g)} + 2NaOH_{(aq)}$$

ELECTROLYSIS OF DILUTE AND CONCENTRATED AQUEOUS SOLUTIONS OF METAL HALIDES

Q.9 Discuss the electrolysis dilute and concentrated aqueous solutions of metal halides. *[10216022]*

Electrolysis Dilute and Concentrated Aqueous Solutions of Metal Halides

$$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

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.

Electrochemical Series

Which of ions sodium, hydrogen, chloride or hydroxide 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 shown in table (16.2) possess more tendency to lose electrons as compared to the elements present at the bottom.

Fig 16.4: Electrolysis of Dilute Solution of Halides

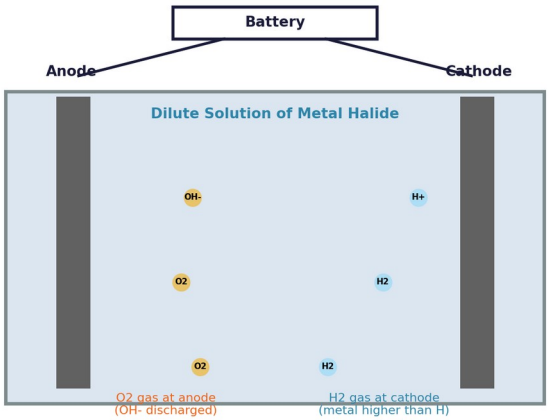

Fig 16.4: Electrolysis of Dilute Solutions of Halides

Following are the categories of metals according to electrochemical series:

- **(i) Metals below hydrogen:** 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.
- **(ii) Metals higher in electrochemical series:** 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.

Interesting Information — S.Q What is the role of electrolysis in industry? [10216023]
Ans. Electrolysis played a crucial role in the discovery and industrial production of elements like fluorine and aluminium.

- **(iii) Lead and nickel metals:** 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.

Reaction at Cathode

$$2M^{+}_{(aq)} + 2e^{-} \rightarrow 2M_{(s)}$$

$$2H^{+}_{(g)} + 2e^{-} \rightarrow H_{2(g)}$$

or

$$2H_2O_{(l)} + 2e^{-} \rightarrow H_{2(g)} + 2OH^{-}_{(aq)}$$

Reaction at Anode

$$2X^{-}_{(g)} \rightarrow X_{2(g)} + 2e^{-}$$

$$4OH^{-}_{(aq)} \rightarrow O_{2(g)} + 2H_2O_{(l)} + 4e^{-}$$

Table 16.2 — Electrochemical Series (Standard Reduction Potentials)

Electrode	Half Reaction	Standard Potential (V)
Lithium	Li+(aq) + e− ⇌ Li(s)	−3.04
Potassium	K+(aq) + e− ⇌ K(s)	−2.93
Calcium	Ca+2(aq) + 2e− ⇌ Ca(s)	−2.86
Sodium	Na+(aq) + e− ⇌ Na(s)	−2.71
Magnesium	Mg+2(aq) + 2e− ⇌ Mg(s)	−2.37
Aluminium	Al+3(aq) + 3e− ⇌ Al(s)	−1.66
Water	2H2O(l) + 2e− ⇌ H2(g) + 2OH−(aq)	−0.82
Zinc	Zn+2(aq) + 2e− ⇌ Zn(s)	−0.76
Iron (II)	Fe+2(aq) + 2e− ⇌ Fe(s)	−0.44
Cadmium	Cd+2(aq) + 2e− ⇌ Cd(s)	−0.40
Nickel	Ni+2(aq) + 2e− ⇌ Ni(s)	−0.25
Lead	Pb+2(aq) + 2e− ⇌ Pb(s)	−0.12

Table 16.2 — Electrochemical Series (Standard Reduction Potentials)		
Hydrogen	$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
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

Stronger reducing agent / Weaker oxidizing agent → (Top of table, e.g. Lithium) | Weaker reducing agent / Stronger oxidizing agent → (Bottom of table, e.g. Fluorine)

ELECTROLYSIS OF CONCENTRATED AQUEOUS SOLUTIONS OF METAL HALIDES

Q.10 Discuss the electrolysis of concentrated solution of metal halides. *[10216024]*

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.

Example: Electrolysis of concentrated aqueous solution of sodium chloride will produce hydrogen gas at the cathode.

Fig 16.5: Electrolysis of Concentrated Copper Chloride

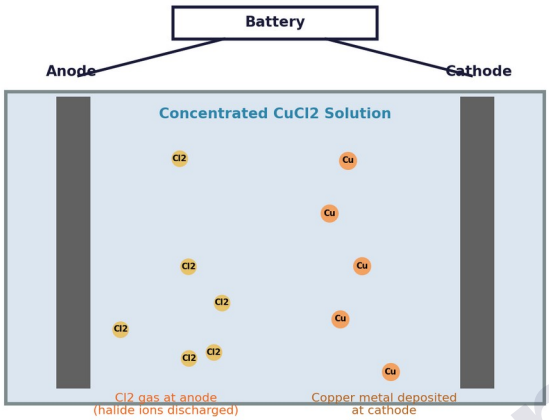

Fig 16.5: Electrolysis of Concentrated Solution of Copper Chloride

Reaction at Cathode

$$2M^{+}_{(aq)} + 2e^{-} \rightarrow 2M_{(s)}$$

$$2H^{+} + 2e^{-} \rightarrow H_2$$

Reaction at Anode

$$2X^{-} \rightarrow X_{2(g)} + 2e^{-}$$

ELECTROLYSIS OF MOLTEN LEAD (II) CHLORIDE

Q.11 How does the electrolysis of molten lead (II) chloride takes place? *[10216025]*

Electrolysis of Molten Lead (II) Chloride

Reaction at Cathode

$$Pb^{2+}_{(l)} + 2e^{-} \rightarrow Pb_{(l)} \quad \text{(Reduction)}$$

Reaction at Anode

$$2Cl^{-}_{(l)} \rightarrow Cl_{2(g)} + 2e^{-} \quad \text{(Oxidation)}$$

Overall Reaction

$$PbCl_{2(l)} \rightarrow Pb_{(l)} + Cl_{2(g)}$$

ELECTROLYSIS OF DILUTE SOLUTION OF SULPHURIC ACID USING INERT ELECTRODES (CARBON OR PT)

Q.12 Why electrolysis of dilute solution of sulphuric acid take place using inert electrodes? *[10216026]*

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 cathode:** Hydrogen ions gain electrons to give hydrogen gas.
- **At 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.
- **Increase in concentration of sulphuric acid:** As the electrolysis moves ahead, the concentration of sulphuric acid increases because the water present is being decomposed.

Reaction at the Cathode

$$2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)} \quad \text{(Reduction)}$$

Reaction at the Anode

$$4OH^-_{(aq)} \rightarrow O_{2(g)} + 2H_2O_{(l)} + 4e^- \quad \text{(Oxidation)}$$

ELECTROLYSIS OF DILUTE AQUEOUS SOLUTION OF COPPER SULPHATE USING INERT ELECTRODES OR COPPER ELECTRODES

Q.13 Explain why the blue colour of copper (II) sulphate fades away during the electrolysis of its aqueous solution. *[10216027]*

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 Cathode

$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)} \quad \text{(Reduction)}$$

Reaction at Anode

$$4OH^-_{(aq)} \rightarrow O_{2(g)} + 2H_2O_{(l)} + 4e^- \quad \text{(Oxidation)}$$

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

Use of Copper Electrodes

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 Cathode

$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)} \quad \text{(Reduction)}$$

Reaction at Anode

$$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e \quad \text{(Oxidation)}$$

Quick Check 16.4 — Why carbon electrodes are used in the electrolysis of concentrated solution of NaCl? [10216028]
Ans. Inert electrode like carbon (graphite) is used in electrolysis of concentrated solution of NaCl because this electrode does not react with the electrolyte or the products formed (Cl2, H2, NaOH).

FUEL CELLS

Q.14 What are fuel cells? Discuss their construction and working to generate electricity. *[10216029]*

Fuel Cells

The devices that generate electricity from a chemical reaction — most commonly between hydrogen and oxygen without burning fuel are called fuel cells.

Construction

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.

Fig 16.6: Hydrogen-Oxygen Fuel Cell

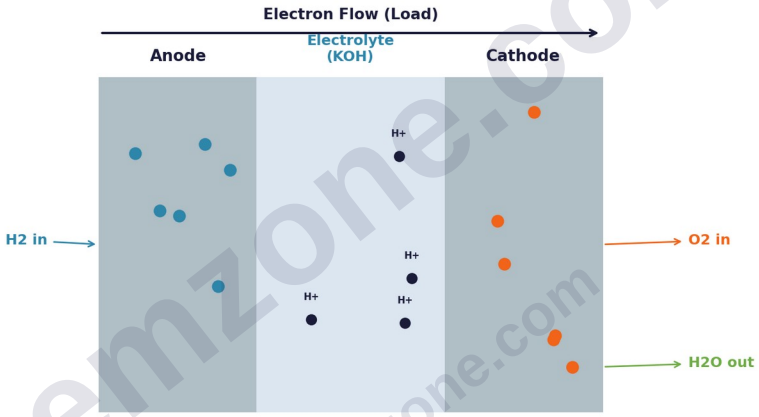

Fig 16.6: Fuel Cell

Working

At the anode, hydrogen is oxidized to water and at the cathode oxygen is reduced to hydroxide ions.

At Anode

$$[H_{2(g)} + 2OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + 2e^-] \times 2$$

At Cathode

$$O_{2(g)} + 2H_2O_{(l)} + 4e^- \rightarrow 4OH^-$$

Overall Reaction

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$$

Products: 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 the advantages and disadvantage of fuel cell in detail. *[10216030]*

Advantages of Fuel Cell

- **(i) Advantages over fossil fuels:** 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) More efficient:** Fuel cells can convert a large proportion about 75% 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) Less noise pollution:** Fuel cells produce less noise pollution.
- **(iv) More sustainable:** 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 Cell

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

Quick Check 16.5 — Q. Which other electrolyte can be used in hydrogen oxygen fuel cell? [10216031]
Ans. Sodium hydroxide (NaOH) and potassium hydroxide (KOH) are commonly used electrolyte in fuel cell. But phosphoric acid, molten carbonates etc can be used as electrolytes.

CORROSION

Q.16 What is corrosion? Explain the reactions involved in rusting of iron. *[10216032]*

Corrosion

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

Oxidation: Iron loses electrons to form Fe (II) ions

$$2Fe_{(s)} \rightarrow 2Fe^{2+}_{(aq)} + 4e^-$$

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

$$O_{2(g)} + 2H_2O_{(aq)} + 4e^- \rightarrow 4OH^-_{(aq)}$$

Conversion of Fe^{2+} to Fe^{3+}

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

$$Fe^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Fe(OH)_{2(s)}$$

$$4Fe(OH)_{2(s)} + O_{2(g)} + 2H_2O_{(l)} \rightarrow 4Fe(OH)_{3(s)}$$

Q.17 Discuss the various methods of prevention of corrosion. *[10216033]*

Prevention of Corrosion

- **(i) 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.

- **(ii) 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.

(iii) 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.

Quick Check 16.6

1. What conditions are required for rusting of iron? *[10216034]*

Ans. Air and moisture are conditions required for frequent rusting of iron.

2. If a sample of iron and a sample of zinc come into contact, which will corrode? Explain your answer. *[10216035]*

Ans. If iron (Fe) and Zinc (Zn) come into to contact with each other, Zn will corrode but Fe remains intact, it is called sacrificial corrosion. Zn is present at the higher position in electrochemical series and have more tendency to get oxidized as compared to iron (Fe). Zn is more active metal than Fe so it gets corroded.

ELECTROPLATING

Q.18 What is electroplating? Give main objectives of electroplating. *[10216036]*

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.

Q.19 Explain the procedure of electroplating in detail. *[10216037]*

Procedure of Electroplating

Electroplating procedure involves the following steps.

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

Fig 16.7: Electroplating

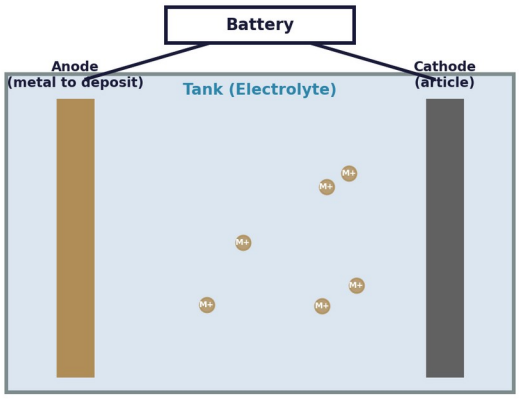

Fig 16.7: Electroplating

Properties of Electrolyte

The electrolyte chosen should have the following properties:

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

Things to Know — S.Q What are noble metals? [10216037a]
Ans. A group of chemically unreactive and valuable metals are called noble metals. They include gold, silver and platinum.

Conditions

The following conditions should be fulfilled to get good deposits.

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

Examples: Silver plating, copper plating etc.

Q.20 Write note on silver and copper plating. *[10216038]*

(i) Silver Plating

Electrodes are dipped in the electrolytic tank containing the electrolyte potassium argento cyanide $K[Ag(CN)_2]$, and electric current is passed. A thin layer of the silver is deposited on the article to be electroplated. The article to be electroplated is made as cathode. Anode is of pure silver.

The following ions are formed in the electrolytic solution.

$$K[Ag(CN)_2]_{(aq)} \rightleftharpoons K^+_{(aq)} + [Ag(CN)_2]^-_{(aq)}$$

$$[Ag(CN)_2]^-_{(aq)} \rightleftharpoons Ag^+_{(aq)} + 2CN^-_{(aq)}$$

Working

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, $K[Ag(CN)_2]$.

(ii) 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

$$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-$$

At Cathode

$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$$

Interesting Information — Q. How does iron undergo chromium plating? [10216039]
Ans. 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.

Quick Check 16.7
1. What is the principle behind electroplating? *[10216040]*
Ans. Electroplating is based on the principle of electrolytic reduction using a direct current to deposit a layer of metal into another metal from an electrolyte solution.
2. How electroplating is used to repair worn-out areas of metal objects? *[10216041]*
Ans. Worn-out or damaged part of metal objects is restored by electroplating using electrolysis to deposit a fresh layer of a metal onto a damaged part. The worn parts act as cathode and metal ions are deposited on it to make smooth and durable coating of metal. In this way, the damaged part of metal objects can be restored.

GALVANIC CELLS

Q.21 What are Galvanic cells? Which type of chemical reactions are taking place in this type of cells? *[10216042]*

Galvanic Cells

The cells in which chemical energy is converted into electrical energy is called galvanic cells.

Explanation

The oxidation-reduction reactions take place at the electrodes when fused ionic compounds or the aqueous solutions of ionic compounds are electrolyzed. The reverse process may also occur. In other words, oxidation-reduction reactions can be used to produce electricity. This can be understood with a very simple process.

Process in Galvanic Cell

- Take a plate of zinc metal and dip it in an aqueous solution of copper sulphate.
- After some time, a pink layer of copper metal will appear on the zinc plate and the blue colour of the solution will become lighter.
- If the quantity of zinc is enough and sufficient time is allowed, the solution will become colourless. In this process, electrons are transferred from zinc to copper ions.
- Every copper ion gets two electrons from zinc atom to get deposited as a metallic copper atom.
- A zinc atom goes into the aqueous solution as Zn^{2+}.

The reaction can be shown as follows:

$$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$$

In other words, zinc is oxidized to $Zn^{2+}(aq)$ while $Cu^{2+}(aq)$ is reduced to copper metal. This reaction can be made a basis for producing electric current in a galvanic cell which is named as Daniel cell.

Q.22 Write construction, working, and chemical reactions of Daniel Cell. *[10216043]*

Daniel Cell

The cell which involves a spontaneous chemical reaction to generate electricity is called a galvanic cell or Daniel cell.

The cell is named after the name of English scientist F. Daniel.

Fig 16.8: Daniel Cell (Zn-Cu Galvanic Cell)

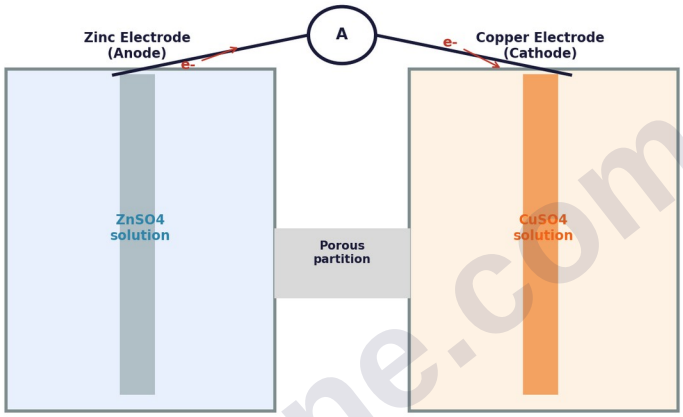

Fig 16.8: Daniel Cell

Construction

- In this type of galvanic cell, a zinc plate is dipped in 1M zinc sulphate solution.
- A copper plate is dipped in 1M copper sulphate solution. The set up makes two half-cells in an electrochemical cell.

Working

- 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.

Chemical Reactions

The following reactions take place at the two electrodes.

$$\text{At anode: } Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^- \quad \text{(Oxidation)} \quad \text{Standard oxidation potential} = +0.76\ \text{V}$$

$$\text{At cathode: } Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)} \quad \text{(Reduction)} \quad \text{Standard reduction potential} = +0.34\ \text{V}$$

$$\text{Overall cell reaction: } Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)} \quad \textbf{Cell voltage } \mathbf{E^o = 1.10\ volts}$$

- 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.
- Daniel cell produces a voltage of around 1.10 volts under standard conditions.

Quick Check 16.8

1. Compare an electrolytic and a galvanic cell. [10216044]

Ans.

Electrolytic Cell	Galvanic Cell
Electrolytic cell: The cell which converts electrical energy into chemical energy is called electrolytic cell. Examples: Down's cell, Nelson's cell etc.	Galvanic cell: The cell which converts chemical energy into electrical energy is called galvanic cell. Examples: Daniel cell, Cu-Zn cell etc.

2. Write down the reactions which occur in a Zn-Mg galvanic cell. [10216045]

Ans. Zn-Mg Galvanic cell.

At cathode: $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ (Reduction)

At anode: $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$ (oxidation)

As Mg is more reactive metal than Zn. So Mg behaves as anode and Zn behaves as cathode. Zn gets reduced and Mg is oxidized.

ORDER OF REACTIVITY OF METALS USING VOLTAGE DATA

Q.23 Discuss the reactivity of Metals using voltage data. Which series is established from this kind of observations? *[10216046]*

Order of Reactivity of Metals using Voltage Data

- According to the available data the reactivity of metals depends upon how readily it is oxidized or lose electrons to form positive ions.
- If a galvanic cell is established between two metals, the more reactive metal will lose electrons and form positive ions.
- The voltage generated by this cell measures the comparative tendencies of the metals to lose electrons.
- A higher voltage indicates that the difference in the tendency of one metal (more reactive metal) to lose electron is far more than the other.

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

MULTIPLE CHOICE QUESTIONS (EXERCISE)

A. Choose and tick the correct answer from the given choices.

1. The oxidation number of Mn in K2MnO4: *[10216047]* (a) +7 (b) +5 (c) +6 (d) −6

2. Which element in the following reaction is being reduced? $4HBr(aq) + MnO2(s) \rightarrow MnBr2(aq) + Br2(g) + 2H2O(l)$ *[10216048]*
(a) Br (b) Mn (c) H (d) Both Mn and H

3. During electrolysis of dilute solution of H2SO4, using Pt electrodes, which process takes place at the cathode? *[10216049]* (a) Reduction (b) Neither oxidation nor reduction (c) Oxidation (d) First oxidation and then reduction

4. Which product you will get at the anode when an aqueous solution of CuSO4 is electrolysed using copper electrodes? *[10216050]*
(a) Cu (b) O2 (c) H2 (d) No product, instead copper metal will go into the solution as Cu2+

5. Which of the following statements is NOT correct about Zn-Cu electrochemical cell? *[10216051]* (a) Anode is negatively charged (b) Reduction occurs at anode (c) Cathode is positively charged (d) Reduction occurs at cathode

6. Which product will be obtained at the anode when a concentrated solution of NaCl will be electrolysed? *[10216052]* (a) Cl2 (b) O2 (c) More O2 and less Cl2 (d) More Cl2 and less O2

7. When a dilute aqueous solution of ZnCl2 is electrolysed which product you will get at the cathode? *[10216053]* (a) H2 (b) Zn (c) O2 (d) Zn and O2

8. What is produced at the cathode in hydrogen-oxygen fuel cell? *[10216054]* (a) H+ (b) O2− (c) H2O (d) OH−

9. Which of the following metals is highest in the electrochemical series? *[10216055]* (a) Tin (b) Iron (c) Magnesium (d) Zinc

SLO Based Multiple Choice Questions

Oxidation Number of atoms in Compounds and Ions

10. Oxidation number of atoms in a compound helps to identify. *[10216056]* (a) Oxidizing agents (b) Reactive agents (c) Reducing agents (d) Both (a) and (c)

11. _____ is an apparent charge on an atom in a compound or anion. *[10216057]* (a) Valency (b) Oxidation number (c) Group number (d) Atomic number

12. The oxidation state of sodium in sodium metal is: *[10216059]* (a) +3 (b) +1 (c) +2 (d) 0

13. Elements of groups 2 shows which of oxidation numbers. *[10216060]* (a) −2 (b) +2 (c) +1 (d) both (a) and (b)

14. The elements of group 17 are normally given oxidation number. *[10216061]* (a) −1 (b) +7 (c) +1 (d) 0

15. In CHCl3, the oxidation number of chlorine is: *[10216062]* (a) +1 (b) +4 (c) −1 (d) −2

16. The oxidation number of O in OF2 is: *[10216063]* (a) +2 (b) −2 (c) +1 (d) −1

17. The sum of oxidation numbers of all the atoms present in a neutral compound is always. *[10216064]* (a) +1 (b) −1 (c) +2 (d) zero

Writing Chemical Formulas of Ionic Compounds by Applying Oxidation Numbers

18. The element calcium is present in group 2 of the periodic table and its principle oxidation number is____. *[10216065]* (a) +1 (b) +2 (c) +3 (d) −2

19. The formula of calcium chloride will be: *[10216066]* (a) CaCl2 (b) CaCl (c) CaCl3 (d) Ca2Cl

Writing Chemical Formula of Ionic Compounds using Ionic Charges

20. MnO4−1 ion has a name: *[10216067]* (a) Manganite ion (b) Permanganes ion (c) Manganese ion (d) Permanganate ion

21. By combining ammonium ion (NH4+1) and sulphate ion (SO4 2−), the compounds formed is: *[10216068]* (a) NH4SO3 (b) (NH4)2(SO4)3 (c) NH4SO4 (d) (NH4)2SO4

22. In writing chemical formula of ionic compounds, the oxidation number of cation will go to _____ of anion. *[10216069]* (a) Subscript (b) Superscript (c) Front (d) Behind

23. The chemical formula of potassium acetate: *[10216070]* (a) CH3COOH (b) CH3COK (c) CH3COOK (d) KCH3OH

24. When calcium (Ca+2) and phosphate (PO4−3) ions reacts, the chemical formula of compound is: *[10216071]* (a) Ca3(PO4)2 (b) Ca2(PO4) (c) Ca4(PO4) (d) Ca2(PO4)3

Oxidation and Reduction

25. Oxidation and reduction are simultaneous process that always occur _____. *[10216072]* (a) Separately (b) Slowly (c) Explosively (d) Together

26. Rusting of iron and burning of magnesium in air are _____ processes. *[10216073]* (a) Simultaneous (b) Reduction (c) Oxidation (d) All of these

27. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ in this reaction, CH4 is being ____ *[10216074]* (a) Reduced (b) Formed (c) Separated (d) Oxidized

28. The reactant which oxidized the other reactant in a chemical reaction is called: *[10216075]* (a) Reducing agent (b) Oxidizing agent (c) Both (a) and (b) (d) Stable agent

29. _____ is a process in which oxidation number of an element increases: *[10216076]* (a) Reduction (b) Decomposition (c) Boiling (d) Oxidation

30. $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$. In this reaction, CuO is called ____. *[10216077]* (a) Oxidizing agent (b) Stable agent (c) Reducing agent (d) All of these

Oxidation Reduction Reactions

31. _____ is a reaction in which simultaneous oxidation and reduction takes place. *[10216078]* (a) Combination (b) Decomposition (c) Neutralization (d) Redox

32. Potassium permanganate acts as a powerful oxidizing agent in _____ medium. *[10216079]* (a) Acidic (b) Neutral (c) Stable (d) All of these

33. Oxalic acid is oxidizes to carbon dioxide by _____. *[10216080]* (a) $KMnO_3$ (b) $KMnO_4$ (c) MnO_2 (d) KOH

34. The oxidation number of carbon (C) in $C_6H_{12}O_6$ is: *[10216081]* (a) +4 (b) −4 (c) zero (d) +2

35. When oxidation number of Mn decreased from +7 to +2, it _____ electrons: *[10216082]* (a) Donates five (b) Accepts five (c) Donates two (d) Accepts two

36. Aqueous potassium iodide is oxidized to iodine in acidic _____ *[10216083]* (a) Potassium mangnate (b) Potassium manganese (c) Potassium permanganate (d) Potassium permanganese

37. The oxidation number of Mn in $(MnO_4)^{-1}$ is: *[10216084]* (a) +2 (b) +5 (c) +6 (d) +7

38. Identify the reducing agent in this reaction: $H_2S + Cl_2 \rightarrow 2HCl + S$ *[10216085]* (a) H_2S (b) HCl_2 (c) HCl (d) S

Electrolysis

39. The breakdown of fused ionic compound or aqueous solution by passing electrify is: *[10216086]* (a) Electrolyte (b) Electrolysis (c) Substitution (d) Electroplating

40. In the process of electrolysis, _____ reaction takes place by using electrical energy. *[10216087]* (a) Spontaneous (b) Slow (c) Formation (d) Non-spontaneous

41. The electrodes in an electrolytic cell are made up of: *[10216088]* (a) Diamond and any other metal (b) Graphite and any other non-metal (c) Steel and any other non-metal (d) Graphite and any other metal

42. In electrolytic cell, the external voltage source forces the reaction to occur by pumping _____ towards one of the electrode: *[10216089]* (a) Hydrogen (b) Proton (c) Electrons (d) Dust

43. The positive ions of the electrolyte are attracted towards the negative electrode and get _____ in electrolytic cell: *[10216090]* (a) Reduced (b) Oxidized (c) Decomposed (d) Unstablized

44. If oxidation and reduction at the electrodes _____, the flow of electricity in the external circuit also _____ in electrolytic cell. *[10216091]* (a) Starts, stops (b) Ceases, stops (c) Ceases, starts (d) Starts, ceases

45. _____ is always occurs at cathode. *[10216092]* (a) Oxidation (b) Reduction (c) Both (a) and (b) (d) Decomposition

Electrolysis of a Concentrated Aqueous Solution of Sodium Chloride (Brine) using Inert Electrode (Graphite/Pt)

46. _____ is formed at anode in electrolysis of a concentrated solution of sodium chloride using inert electrode. *[10216093]* (a) NaOH (b) Cl_2 (c) H_2 (d) O_2

47. In electrolysis of concentrated aqueous solution of brine, anode must be made up of graphite to avoid reaction with: *[10216094]* (a) Chlorine gas (b) Hydrogen gas (c) HNaOH (d) Oxygen gas

48. _____ both migrate towards the cathode in the electrolysis of concentrated aqueous solution of sodium chloride. *[10216095]* (a) Cl^- and H_2O (b) Na^+ and H_2O (c) OH^- and Na^+ (d) H^+ and Cl^-

49. Hydrogen ions have a _____ tendency to get reduced as composed to sodium ions (Na^{+1}). *[10216096]* (a) Greater (b) Lesser (c) Equal (d) Almost equal

50. The solution around the cathode becomes _____ due to formation of OH^{-1} ions in electrolysis of concentrated aqueous solution of sodium chloride. *[10216097]* (a) Acidic (b) Neutral (c) Stable (d) Alkaline

51. In electrolysis of concentrated aqueous solution of sodium chloride, the volumes of hydrogen and chlorine gases obtained are: *[10216098]* (a) Unequal (b) $Cl_2 > H_2$ (c) Equal (d) $H_2 > Cl_2$

Electrolysis of Dilute and Concentrated Aqueous Solution of Metal Halides

52. Water is a _____ electrolyte. *[10216099]* (a) Strong (b) Very strong (c) Moderate (d) Weak

53. If aqueous NaCl is electrolyzed, sodium (Na+1) and _____ ions are attracted towards cathode. *[10216100]* (a) H+1 (b) Cl−1 (c) OH−1 (d) All of these

54. The elements at the top of electrochemical series possess more tendency to _____. *[10216101]* (a) Gain electrons (b) Lose electrons (c) Stabilize (d) Share electrons

55. If dilute copper chloride (CuCl2) is electrolyzed the product at cathode: *[10216102]* (a) Cu(s) (b) Cl2(g) (c) H2(g) (d) O2(g)

56. _____ is obtained at cathode, if dilute magnesium chloride is electrolyzed. *[10216103]* (a) H2(g) (b) Cl2(g) (c) Mg(s) (d) O2(g)

57. More complicated picture is given by electrolysis of dilute and concentrated halides solutions of metals: *[10216104]* (a) Nickel and sodium (b) Lead and sodium (c) Lead and nickel (d) Copper and silver

58. At anode, the electrolysis of dilute solutions of halides will deposit: *[10216105]* (a) Hydrogen (b) Hydroxide ion (c) Halogen (d) Oxygen

59. Electrolysis played an important role in industrial preparation of element: *[10216106]* (a) Carbon (b) Nitrogen (c) Aluminum (d) Uranium

60. The element with lowest standard reduction potential in electrochemical series: *[10216107]* (a) H2 (b) Li (c) F2 (d) Au

Electrolysis of Concentrated Aqueous Solutions of Metal Halides

61. In electrolysis of concentrated aqueous solution of metal halides, H+1 ions are discharged at cathode, if metal is present higher than _____ in electrochemical series. *[10216108]* (a) Sodium (b) Magnesium (c) Hydrogen (d) Copper

62. The electrolysis of concentrated aqueous solution of sodium chloride produces _____ at the cathode. *[10216109]* (a) Chlorine gas (b) Sodium metal (c) Hydrogen gas (d) Oxygen gas

63. In the electrolysis of concentrated aqueous solution of copper chloride _____ is discharged at the cathode. *[10216110]* (a) Cl2 (b) H2 (c) Ag (d) Cu

Electrolysis of Molten Lead (II) Chloride

64. Lead chloride is melted upto: *[10216111]* (a) 601°C (b) 501°C (c) 701°C (d) 401°C

65. In the electrolysis of molten lead chloride, two insert electrodes immersed in molten salt are: *[10216112]* (a) Graphite or iron (b) Iron or copper (c) Graphite or platinum (d) Graphite or copper

66. In the electrolysis of molten lead (II) chloride, the product formed at cathode: *[10216113]* (a) Chlorine gas (b) O2 gas (c) H2 gas (d) Lead metal

Electrolysis of Dilute Solution of Sulphuric Acid using Inert Electrodes (Carbon and Pt)

67. Electrolysis of dilute aqueous solution of sulphuric acid produces: *[10216114]* (a) H2 gas (b) H2 and O2 gas (c) O2 gas (d) SO_4^{2-} ions

68. Why inert electrodes used in electrolysis of dilute solution of sulphuric acid? *[10216115]* (a) Because they did not react with electrolyte (b) Because they did not react with water (c) Because they only react with electrolyte (d) Because they react with water and electrolyte

69. The correct reaction in electrolysis of dilute sulphuric acid at cathode is: *[10216116]* (a) $H_2(g) \rightarrow 2H^+(aq) + 2e^-$ (b) $4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$ (c) $H_2(g) \rightarrow H_2(l)$ (d) $2H^+(aq) + 2e^- \rightarrow H_2(g)$

Electrolysis of Dilute Aqueous Solutions of Copper Sulphate using Inert Electrodes or Copper Electrode

70. In the electrolysis of aqueous solution of copper sulphate using inert electrodes the product formed at cathode: *[10216117]* (a) O2 (b) Cu (c) H2 (d) H2 and O2

71. The concentration of _____ decreases slowly and blue colour of the solution fades away with the passage of time in the electrolysis of aqueous solution of copper sulphate. *[10216118]* (a) SO_4^{2-} ions (b) Fe^{+2} ions (c) Cu^{+1} ions (d) OH^- ions

72. The products formed at cathode and anode in electrolysis of dilute aqueous solution of copper sulphate using inert electrode: *[10216119]* (a) Cu(s) and H2(g) (b) H2(g) and O2(g) (c) Fe(s) and O2(g) (d) Cu(s) and O2(g)

Fuel Cells

73. Fuel cells converts chemical energy into _____ *[10216120]* (a) Vibrational energy (b) Electrical energy (c) Rotational energy (d) Translational energy

74. The gaseous fuel used in fuel cell to produce electricity are: *[10216121]* (a) H2 and O2 (b) H2 and N (c) N2 and Cl2 (d) CO2 and O2

75. Electrodes are made of porous compressed carbon saturated with platinum act as _____ in fuel cell: *[10216122]* (a) Reactant (b) Product (c) Activator (d) Catalyst

76. The overall reaction in fuel cell produces: *[10216123]* (a) O2 (b) H2O (c) KOH (d) H2

77. Fuel cell converts _____% of bond energy into electricity: *[10216124]* (a) 20 (b) 40 (c) 75 (d) 95

78. Fuel cells produces less _____ pollution. *[10216125]* (a) Air (b) Water (c) Solid (d) Noise

79. Hydrogen used in fuel cells can be obtained from _____. *[10216126]* (a) Non-renewable sources (b) Renewable sources (c) Natural sources (d) All of these

80. Fuel cells and electric motors are less _____ than petrol engines. *[10216127]* (a) Stable (b) Effective (c) Durable (d) Compact

Corrosion

81. Corrosion is a _____ process due to which a metal suffers gradual decay when it interacts chemically with its environment. *[10216128]* (a) Artificial (b) Fast (c) Moderate (d) Natural

82. Galvanization is a common method to protect iron or steel from: *[10216129]* (a) Rusting (b) Burning (c) Neutralization (d) Substitution

83. In sacrificial protection a more reactive metal like zinc or magnesium is corrode in place of less reactive metal like: *[10216130]* (a) Lithium (b) Sodium (c) Potassium (d) Iron

84. The difference in the activity of two metals (Mg and Fe) cause a current to glow, producing corrosion on: *[10216131]* (a) Less metal (b) Stable metal (c) More active metal (d) Moderate active metal

Electroplating

85. The main objectives of electroplating are: *[10216132]* (a) Decoration (b) Repair (c) Protection (d) All of these

86. The article to be electroplated is made: *[10216133]* (a) Anode (b) Electrolyte (c) Cathode (d) Solution

87. Electroplating is carried out in a tank made of: *[10216134]* (a) Paper (b) Plastic (c) Rubber (d) Cement

88. The condition when one metal is deposited on another to get good deposits: *[10216135]* (a) High current density (b) Low current density (c) High temperature (d) High pH

89. In silver plating, the electrolyte used: *[10216136]* (a) K[Ag(Cl)2] (b) K[Ag(H2O)2] (c) K[Fe(CN)2] (d) K[Ag(CN)2]

90. _____ is used as an anode in electroplating of copper. *[10216137]* (a) CuSO4 (b) Silver metal (c) Iron metal (d) Copper metal

Galvanic Cell

91. Oxidation-reduction reaction can be used to produce electricity in _____ cell. *[10216138]* (a) Electrolytic (b) Galvanic (c) Down's (d) Nelson

92. In a galvanic cell, electrons are transferred from: *[10216139]* (a) Zinc to copper (b) Copper to zinc (c) Copper to iron (d) Zinc to sodium

93. In galvanic cell, two solutions are kept separate by: *[10216140]* (a) Iron (b) Porous partition (c) Non-porous partition (d) Glass wall

94. Daniel cell produces a voltage of around _____ volts under standard conditions: *[10216141]* (a) 0.110 (b) 0.01 (c) 1.10 (d) 0.76

Order of Reactivity of Metals using Voltage Data

95. More easily a metal can oxidize or lose electron, more it will be: *[10216142]* (a) Reactive (b) Inert (c) Stable (d) Strong

96. A galvanic cell of magnesium and copper produces more voltage than a cell of copper and: *[10216143]* (a) Hydrogen (b) Zinc (c) Nickel (d) Iron

97. The reactivity series of metals is called: *[10216144]* (a) Electromagnetic series (b) Electrochemical series (c) Chemical series (d) Stability series

Answer Key

1	2	3	4	5	6	7
c	b	a	d	b	a	a
8	**9**	**10**	**11**	**12**	**13**	**14**
d	c	d	b	d	d	a
15	**16**	**17**	**18**	**19**	**20**	**21**
c	a	d	b	a	d	d
22	**23**	**24**	**25**	**26**	**27**	**28**
c	c	a	d	c	d	b
29	**30**	**31**	**32**	**33**	**34**	**35**
d	a	d	a	b	c	b
36	**37**	**38**	**39**	**40**	**41**	**42**
c	d	a	b	d	d	c
43	**44**	**45**	**46**	**47**	**48**	**49**
a	b	b	b	a	b	a
50	**51**	**52**	**53**	**54**	**55**	**56**
d	c	d	b	b	a	a
57	**58**	**59**	**60**	**61**	**62**	**63**
c	d	c	b	c	c	d
64	**65**	**66**	**67**	**68**	**69**	**70**
b	c	d	b	a	d	b
71	**72**	**73**	**74**	**75**	**76**	**77**
a	d	b	a	d	b	c
78	**79**	**80**	**81**	**82**	**83**	**84**
d	b	c	d	a	d	c
85	**86**	**87**	**88**	**89**	**90**	**91**
d	c	d	a	d	d	b
92	**93**	**94**	**95**	**96**	**97**	
a	b	c	a	d	b	

SHORT ANSWER QUESTIONS (EXERCISE)

B. Write Short Answers.

1. What happens when electricity is passed through an aqueous solution of NaCl? *[10216145]*

Ans. When electricity is passed through aqueous solution of sodium chloride both sodium and hydrogen cations are attracted towards the cathode. As sodium lies above hydrogen in electrochemical series, hydrogen ion is discharged at cathode and chlorine gas is produced at anode. In the solution, sodium ions Na+ combine with OH−. Therefore, it will become alkaline due to formation of NaOH. Main products: Hydrogen gas at cathode; Chlorine gas at anode; Sodium hydroxide NaOH in solution.

2. What are the main objectives of electroplating? *[10216146]*

Ans. 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.

3. Mention one difference between an electrolytic cell and a galvanic cell. *[10216147]*

Ans.

Electrolytic cell	Galvanic cell
A cell which converts electrical energy into chemical energy is called electrolytic cell. Examples: Electrolysis of molten NaCl (Down's cell)	A cell which uses chemical energy to produce electric current is called Galvanic cell. Examples: Daniel cell, batteries etc.

4. Through which electrode, electrons leave the electrolytic cell? *[10216148]*

Ans. It is a negative electrode that is cathode through which electrons leave the external circuit and move towards positive electrode (anode).

5. How does an electrolyte conduct electricity? *[10216149]*

Ans. Electrolytes conduct electricity through the movement of ions. When dissolved in water or melted electrolyte breaks into cation and anions (positive and negative ions). Cations move towards the cathode and anions move towards the anode. This movement of ions is the reason of flow of electric current.

6. What will happen if a strip of Cu metal is dipped in FeSO4 solution? *[10216150]*

Ans. In the electrochemical series, copper is below iron, so it can't displace iron from FeSO4 solution. So, no reaction occurs because copper is less reactive than iron. The copper strip remains unchanged and the green color of FeSO4 solution stays the same.

7. Why fuel cells are regarded as environment friendly? *[10216151]*

Ans. Fuel cells are environmental friendly because: (i) Fuel cell 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) Less noise pollution: Fuel cells produce less noise pollution. (iii) More sustainable: Hydrogen used in fuel cells can be obtained from renewable sources like solar or wind power.

SLO BASED SHORT ANSWER QUESTIONS

Oxidation Number of Atoms in Compounds and Ions

8. What is oxidation number? *[10216152]*

Ans. The oxidation number is an apparent charge of an atom in a compound or an ion. It may have a positive or negative ion. Example: Oxidation number O in H2O is −2.

9. Write down advantages of oxidation number. *[10216153]*

Ans. Oxidation number of atoms in compounds or ions helps to identify either the reaction is oxidation or reduction. Oxidation numbers are also used to write chemical formulas and names of chemical compounds.

10. Find out the oxidation number of Cr in Cr2O32−. *[10216154]*

Ans. Solution

$$2(Cr) + 3(-2) = -2$$
$$2(Cr) + (-6) = -2$$
$$2(Cr) = -2 + 6$$
$$\mathbf{2Cr/2 = 4/2 \rightarrow Cr = +2}$$

Writing Chemical Formulas of Ionic Compounds using Ionic Charges

11. How chemical formula of ionic compound containing polyatomic ions can be written? *[10216155]*

Ans. When writing the formulae of ionic compound containing more than one polyatomic ions the chemical formula of ions is taken in brackets and indicate the number of such ions with subscript. Example: The formula of ammonium carbonate is (NH4)2CO3.

12. Write down the formula of magnesium nitrate. *[10216156]*

Ans. The formula of magnesium nitrate is Mg(NO3)2.

Oxidation and Reduction

13. Differentiate between oxidation and reduction. *[10216157]*

Ans.

Oxidation	Reduction
Oxidation is defined as a process in which oxygen is either added or hydrogen is removed. Example: Oxygen combine with other elements in the formation of their oxides. 4Fe + 3O2 → 2Fe2O3	Reduction is defined as a process in which either addition of hydrogen or removal of oxygen takes place. Example: Hydrogen is added in chlorine to give hydrogen chloride gas. H2 + Cl2 → 2HCl

14. Which process show a decrease in oxidation number of an element? *[10216158]*

Ans. Reduction is a process in which oxidation number of an element decreases. Example: Mn7+ + 5e− → Mn2+. In this example oxidation number of manganese is reduced from +7 [Mn(VII)] to +2 [Mn(II)].

15. What are reducing agents? Give example. *[10216159]*

Ans. Reducing Agents: The species which oxidizes other species in a chemical reaction are called reducing agent. Example: Hydrogen is removed from a compound in the following reaction. CH4 + 2O2 → 2CO2 + 2H2O. CH4 is being oxidized while O2 is being reduced. CH4 is reducing agent in this reaction.

Oxidation Reduction Reactions

16. What are redox reactions? *[10216160]*

Ans. Redox reactions: A chemical reaction which involves simultaneous oxidation and reduction is called oxidation-reduction or redox reaction. Example: H2S + Cl2 → 2HCl + S

17. How potassium permanganate changes its colour in reaction with oxalic acid? *[10216161]*

Ans. Potassium permanganate oxidizes oxalic acid into carbon dioxide. During the reaction dark purple colour of potassium permanganate changes to colourless.

Electrolysis

18. Define electrolysis. *[10216162]*

Ans. Electrolysis: When electricity is passed through fused ionic compounds or their aqueous solutions, they are decomposed into their constituents. Such compounds are called electrolytes and the process is called electrolysis.

19. How non-spontaneous reaction is carried out in electrolysis? *[10216163]*

Ans. In the process of electrolysis, the non-spontaneous reaction takes place with the help of electrical energy. The external voltage source forces the non-spontaneous reaction by pumping electrons towards one of the electrodes.

20. What kind of reaction takes place at cathode? *[10216164]*

Ans. The positive ions of the electrolyte are attracted towards the negative electrode that is cathode, and get reduced by accepting electrons from it. So reduction takes place at cathode.

Electrolysis of a Concentrated Aqueous Solution of Sodium Chloride (Brine) using Inert Electrodes (Graphite/Pt)

21. Give products of electrolysis of a concentrated aqueous solution of sodium chloride. *[10216165]*

Ans. 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 as a by-product.

23. Why solution become alkaline around cathode in electrolysis of concentrated aqueous sodium chloride? *[10216167]*

Ans. The OH^{-1} ions are formed around cathode due to decomposition of water. So the solution around the cathode becomes alkaline due to the formation of hydroxide ions.

Electrolysis of Dilute and Concentrated Solution of Metal Halides

24. Why situation becomes complicated in the electrolysis of aqueous solution of metal halides? *[10216168]*

Ans. The situation becomes complicated in the electrolysis of aqueous metal halide due to presence of water. Water also ionizes in small amount to give hydrogen and hydroxide ions.

25. Which kind of elements have more tendency to get oxidized? *[10216169]*

Ans. The elements present at the top of the electrochemical series possess more tendency to lose electron and gets oxidized. Example: Na, Li, K etc have more tendency to get oxidized as compared to other elements.

22. Write the reaction at cathode and anode during electrolysis of brine. *[10216166]*

Ans. At cathode: Na+ and H2O move towards cathode. Water molecules or hydrogen ions pick up electrons to form hydrogen gas.

$$2H_2O_{(l)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}$$

At anode: Cl^- ions move towards anode and gets oxidized by giving electrons.

$$2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$$

Electrolysis of Concentrated Aqueous Solution of Metal Halides

27. What happens with metal at cathode in electrolysis of concentrated aqueous solution of metal halides? *[10216171]*

Ans. If concentrated aqueous solution of metal halide is electrolysed, hydrogen ions are discharged at the cathode, if the metal is present higher than hydrogen in electrochemical series. If metals like silver and copper which are lower than hydrogen in electrochemical, then metal ions are discharged at cathode.

26. Write down reaction at cathode in electrolysis of concentrated solution of metal halide. *[10216170]*

Ans. At cathode: If aqueous solution of metals like nickel or copper is electrolyzed, these are two possibilities. i. If the solution is concentrated, metal is deposited at cathode. ii. If solution is very dilute hydrogen gas is evolved. At in-between concentration both hydrogen and metal are obtained.

$$2M^+_{(aq)} + 2e^- \rightarrow 2M_{(s)}$$
$$2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$$

Electrolysis of Molten Lead (II) Chloride

28. Discuss the reaction at electrodes in electrolysis of concentrated aqueous solution of metal halides. *[10216172]*

Ans. The reactions at the electrodes are:

At cathode: Metal or hydrogen is produced.

$$2M^+_{(aq)} + 2e^- \rightarrow 2M_{(g)}$$
$$2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$$

At anode: Halogen molecules are generally formed.

$$2X^-_{(aq)} \rightarrow X_{2(g)} + 2e^-$$

29. Give the reactions at electrodes when molten lead (II) chloride is electrolyzed. *[10216173]*

Ans. Following are the reactions at electrodes.

At cathode: Molten lead metal is obtained at cathode.

$$Pb^{2+}_{(l)} + 2e^- \rightarrow Pb_{(l)} \quad \text{(oxidation)}$$

At Anode: Chlorine gas is formed at anode.

$$2Cl^-_{(l)} \rightarrow Cl_{2(g)} + 2e^- \quad \text{(reduction)}$$

30. What should be the temperature maintained in electrolysis of molten lead (II) chloride? *[10216174]*

Ans. Electrolyte must be heated upto 501°C to get molten lead chloride. Heat will render ions free to move and conduct electricity. The electrolysis must be carried out at the temperature above the melting point of lead chloride.

Electrolysis of Dilute Solution of Sulphuric Acid using Inert Electrodes (Carbon or Pt)

31. Why inert electrodes are used when dilute aqueous solution of sulphuric acid is electrolysed? *[10216175]*

Ans. Inert electrodes like platinum or graphite are used for the electrolysis of dilute solution of sulphuric acid because they did not react with the electrolyte.

Electrolysis of Dilute Aqueous Solution of Copper Sulphate using Inert Electrodes or Copper Electrodes

32. What are reactions at electrodes in electrolysis of aqueous solution of copper sulphate using inert electrode? *[10216176]*

Ans. The electrolysis of aqueous solution of copper (II) sulphate with inert electrodes results in the decomposition of water and copper (II) sulphate.

At cathode: Copper metal is deposited at cathode (inert electrode).

$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)} \quad \text{(Reduction)}$$

At anode: Oxygen is produced at the anode.

$$4OH^- \rightarrow O_{2(g)} + 2H_2O_{(l)} + 4e^- \quad \text{(oxidation)}$$

33. Give reactions at electrodes in electrolysis of aqueous solution of copper sulphate using copper electrode. *[10216177]*

Ans. In the electrolysis of aqueous solution of copper sulphate using copper electrodes copper metal from copper anode start going into the solution in the form of copper ions and electrode mass starts decreasing. At cathode: Cu2+(aq) + 2e− → Cu(s) (Reduction). At anode: Cu(s) → Cu2+(aq) + 2e− (Oxidation)

Fuel Cell

34. Define fuel cell. *[10216178]*

Ans. Fuel Cells: Fuel cells provide a method in which chemical energy is converted into electrical energy. These cells use gaseous fuels like hydrogen and oxygen to produce electricity.

35. Give reactions at electrodes in fuel cell. *[10216179]*

Ans. At anode: Hydrogen is oxidized to water.

$$[H_{2(g)} + 2OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + 2e^-] \times 2$$

At cathode: Oxygen is reduced to hydroxide ions.

$$O_{2(g)} + 2H_2O_{(l)} + 4e^- \rightarrow 4OH^-_{(aq)}$$

36. Give two advantages of fuel cells. *[10216180]*

Ans. Advantages of fuel cells: i. Fuel cells produce less noise pollution. ii. Hydrogen used in fuel cells can be obtained from renewable sources like solar or wind power. This fact renders fuel cells more sustainable.

37. Give two disadvantages of fuel cells. *[10216181]*

Ans. 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.

Corrosion

38. What is corrosion? *[10216182]*

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

39. How corrosion is prevented by coating of the metal surface? *[10216183]*

Ans. 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.

40. Write a short note on galvanization method for prevention of corrosion. *[10216184]*

Ans. Galvanization: It 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.

41. What is sacrificial protection? *[10216185]*

Ans. 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.

Electroplating

42. Give three main objectives of electroplating. *[10216186]*

Ans. 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.

43. What should be the properties of electrolytes in electroplating? *[10216187]*

Ans. 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

44. Write a short note on working of silver plating. *[10216188]*

Ans. The electrolyte used is the solution of potassium argento cyanide, K[Ag(CN)2], and anode is of pure silver. The following ions are formed in the electrolytic solution.

$$K[Ag(CN)_2]_{(aq)} \rightleftharpoons K^+_{(aq)} + [Ag(CN)_2]^-_{(aq)}$$
$$[Ag(CN)_2]^-_{(aq)} \rightleftharpoons Ag^+_{(aq)} + 2CN^-_{(aq)}$$

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].

45. What is copper plating? *[10216189]*

Ans. In copper plating, the anode is made of copper metal and the electrolyte is CuSO4 to which a few drops of sulphuric acid are added to prevent its hydrolysis. The article to be electroplated is made the cathode. At anode: Cu(s) → Cu2+(aq) + 2e−. At cathode: Cu2+(aq) + 2e− → Cu(s)

Galvanic Cell

46. What is Galvanic cell? Give example. *[10216190]*

Ans. The cell which converts chemical energy into electrical energy is called galvanic cell. Example: Daniel Cell

47. Give reactions at electrodes in Cu-Zn galvanic cell. *[10216191]*

Ans. 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.

Chemical reactions:

$$\text{At anode: } Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^- \text{ (Oxidation)}$$
$$\text{At cathode: } Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)} \text{ (Reduction)}$$
$$\text{Overall: } Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$$

Order of Reactivity of Metals using Voltage Data

48. How reactivity of metal can be predicted by voltage data? *[10216192]*

Ans. According to the available data the reactivity of metals depends upon how readily it is oxidized or lose electrons to form positive ions. If a galvanic cell is established between two metals, the more reactive metal will lose electrons and form positive ions.

49. What happens in Cu-Mg galvanic cell? *[10216193]*

Ans. 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.

CONSTRUCTED RESPONSE QUESTIONS

1. Calculate the oxidation number of chlorine atom in Ca(ClO3)2. *[10216194]*

Ans. Solution

Oxidation number (O.N) of O = -2, Oxidation number of Ca = +2, O.N of Cl = x = ?

O.N of Ca + 2(O.N of Cl) + 6(O.N of O) = 0

+2 + 2x + 6(-2) = 0

2 + 2x - 12 = 0

$$2x = 10$$

$$\mathbf{x = 10/2 = +5 \quad \rightarrow \quad O.N \text{ of } Cl = x = +5}$$

2. How can H2O2 behave as an oxidizing agent? *[10216195]*

Ans. An oxidizing agent is a substance that gives oxygen or accepts electron from another substance. H2O2 behaves as an oxidizing agent because it provides oxygen to other and accepts electron from other compound. In hydrogen peroxide, oxygen has an oxidation state of −1, which is unstable it readily gets reduced to −2 (as in water) causing the other substance to be oxidized.

Example: When it reacts with potassium iodide (KI) in acidic medium:

$$H_2O_2 + 2I^- + 2H^+ \rightarrow I_2 + 2H_2O$$

Iodide ions (I−) are oxidized to iodine (I2) and H2O2 acts as an oxidizing agent.

3. Why different products are formed at the electrodes when a dilute and concentrated aqueous solutions of NaCl are electrolysed? *[10216196]*

Ans. Oxygen gas is produced: The difference in products is due to the fact that ions that get discharged at the electrodes depend on their concentration and ease of discharge. In both dilute and concentrated sodium chloride solutions, these species are present: Na+, Cl−, H+ and OH− (from water).

At cathode (same in both cases): Sodium is higher in electrochemical series, and being highly reactive nature of sodium, water is reduced instead of sodium.

$$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$$

At anode — In dilute solution: Cl− concentration is low. OH− are preferentially discharged.

$$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$$

In concentrated solution: Cl− ion concentration is high so Cl− ions are discharged more easily than OH−. Chlorine gas is produced. So concentration affects which ions are preferentially discharged.

$$2Cl^- \rightarrow Cl_2 + 2e^-$$

5. If a galvanic cell is established between Cu and Ag, what reactions do you expect at the electrodes? *[10216198]*

Ans. In a Cu-Ag galvanic cell, silver is noble metal and has higher reduction potential so it gets reduced.

$$\text{At anode (Oxidation — Cu electrode): } Cu \rightarrow Cu^{2+} + 2e^-$$

$$\text{At cathode (Reduction — Ag electrode): } 2Ag^+ + 2e^- \rightarrow 2Ag$$

4. Why corrosion becomes fast during rainy season? *[10216197]*

Ans. Corrosion becomes fast during rainy season because there is more moisture everywhere. This moisture creates a thin form of water and acts like an electrolyte letting electrons move easily to accelerate corrosion. Rain water contains more dissolved oxygen which also fuels oxidation (rusting in iron and steel). Rain dissolves pollutants like SO2, CO2 and NOx from the air making acid rain. These acids increase corrosion rates especially in industrial areas.

8. How can the chemical reactivity of metals be compared by using electrochemical series? *[10216201]*

Ans. Chemical reactivity of metals can be compared using the electrochemical series by looking at their standard reduction potentials (E° values). It is a series in which metals are arranged in order of their reduction potential.

Metals with more negative E° values have a greater tendency to lose electrons means get oxidized. These metals are more reactive because they readily form positive ions.

When we compare metals, then a metal higher in reactivity series lies higher in electrochemical series more negative E°. While a metal lower in reactivity series lies lower in the electrochemical series with more positive E°.

Example: Zinc E° = −0.76V is more reactive than copper (E° = +0.34V) so zinc can displace copper from its salt solution but copper cannot displace zinc.

6. What is the function of a porous partition in a galvanic cell? *[10216199]*

Ans. In a galvanic cell, the porous partition acts as a separator that allows ions to pass but prevents mixing of the solution. It is also called salt bridge. It performs following functions: It maintains electrical neutrality, by allowing ions to migrate between the two half cells so charge does not build up and stop the reaction. It completes the internal circuit. Prevents direct mixing of electrolytes. Reduces diffusion of solutions.

7. How do fuel cells produce electricity? *[10216200]*

Ans. See Q.14 from theory (Construction and working of fuel cells).

DESCRIPTIVE QUESTIONS (EXERCISE)

1. Explain with examples how the formula of ionic compound can be written using oxidation numbers. *[10216202]*

Ans. See Q.3 from theory.

2. Find the oxidation numbers of the indicated elements in the following compounds: *[10216203]*

$$Al(NO_3)_3,\ H_2\underline{S}_2O_7,\ Na_2\underline{S}_2O_3$$

Ans.

(i) Al(NO3)3

O.N of Al = +3, O.N of O = -2, O.N of N = x

$(+3) + 3x + 9(-2) = 0 \rightarrow 3x - 15 = 0 \rightarrow x = 15/3 = +5$

So O.N of N = +5

(ii) H2S2O7

O.N of H = +1, O.N of O = -2, O.N of S = x

$2(+1) + 2(x) + 7(-2) = 0 \rightarrow 2x - 12 = 0 \rightarrow x = 12/2 = 6$

So O.N of S = +6

(iii) Na2S2O3

O.N of Na = +1, O.N of O = -2, O.N of S = x

$2(+1) + 2(x) + 3(-2) = 0 \rightarrow 2x - 4 = 0 \rightarrow x = 4/2 = +2$

O.N of S = +2

3. Write down the formulae of the following compounds using ionic charges: Sodium hypochlorite, calcium oxalate, ammonium phosphate and calcium nitrite. *[10216204]*

Ans. Given the formulas of ionic compounds using ionic charges.

(i) Sodium hypochlorite: NaClO
(ii) Calcium oxalate: CaC2O4
(iii) Ammonium phosphate: (NH4)3PO4
(iv) Calcium nitrite: Ca(NO2)2

4. Describe the working of a fuel cell. *[10216205]*

Ans. See Q.14 from theory.

5. Discuss the electrolysis of concentrated solutions of metal halides. *[10216206]*

Ans. See Q.10 from theory.

6. Explain electroplating giving examples. *[10216207]*

Ans. See Q.19 from theory.

7. What is corrosion? Explain the reactions involved in corrosion of iron. *[10216208]*

Ans. See Q.16 from theory.