CHAPTER # 12
ELECTROCHEMISTRY

Q1. Define electrochemistry.

Ans: Electrochemistry:

Electrochemistry is the branch of chemistry, which deals with inter-conversion of electrical energy and chemical energy.

Q2. Explain redox reaction.

Ans: Redox reaction:

Those reactions in which gain or loss of electrons occur are called redox reactions. In other words oxidation-reduction reactions are called redox reactions.

Explanation:

Redox reaction involves transfer of electrons from one substance to another. It deals with efficient sources of energy such as batteries, fuel cells etc.

Q3. What do you know about electrochemical processes?

Ans: Electrochemical processes:

Electrochemical processes are redox (oxidation-reduction) reactions in which the energy released by a spontaneous reaction is converted to electricity or in which electrical energy is used to cause a non-spontaneous reaction to occur.

Q4. Explain oxidation-reduction reaction with the help of example.

Ans: Oxidation-Reduction reaction:

Oxidation:

"A reaction in which a substance loses electrons is called **oxidation**.

Reduction:

While the reaction in which a substance gains electrons is called **reduction**.

Redox reaction:

Oxidation – reduction reactions are also known as redox reactions.

Explanation:

When a piece of Zinc metal is dipped in an aqueous solution of $CuSO_4$, it is observed that a dark brown layer of copper begins to form at Zinc surface. At the same time the solution loses its blue colour. If we analyses this solution we find that Zn^{++} ions are present in solution. The change can be described by the following chemical equation.

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

This reaction can be described in terms of two half reactions.

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

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

In this reaction Zinc metal loses two electrons and changes into Zn^{+2} ions while Cu^{+2} ions gain two electrons and give copper metal. The two processes taking place simultaneously are called oxidation – reduction reactions.

Note:

In the redox reaction element undergoing oxidation or reduction undergo a change in their oxidation number

Q5. Define oxidation number and also write the rules through which oxidation number can be assigned.

Ans: Oxidation Number:

The oxidation number or oxidation state, is defined as the apparent change, positive or negative which an element would have in a compound.

Rules:

i. The oxidation state of a free element is always zero e.g., oxidation state of Zn, Na, H in H_2, S in S_8 etc is zero.

ii. In simple ions, oxidation state is same as their charge e.g., oxidation state of Na in Na^{+1} and Ca in Ca^{+2} are +1 and +2 respectively.

iii. In a complex ion the total sum of oxidation states of atoms Is equal to the charge on their ion, e.g., in CO_3^{-2}, the sum of oxidation states of C and 3O atoms is -2. Similarly, in NH_4^{+1} the sum of oxidation states of N and 4H atoms is +1.

iv. The oxidation number of each of the atoms in a molecule or compound counts separately and their algebraic sum is zero e.g., In HCl, the sum of oxidation states of H and Cl atoms is zero. Similarly in CO_2, the sum of oxidation states of one C and 2 oxygen atoms is zero.

v. The more electronegative element has the negative oxidation number.

Elements	Oxidation State
Group-IA	+1
Group-IIA	+2
Group-IIIA	+3
H	+1 (except in metal hydrides where it is -1)
Group-VIIA	-1
O	-2 (except peroxides and in OF_2)

Example 12.1:

Calculate the Oxidation number of Mn in $KMnO_4$.

Solution: Oxidation number of K = +1

Oxidation number of O = − 2

Oxidation number of Mn = x

In compounds, the algebraic sum of the oxidation numbers of all the atoms is zero.

K		Mn		O_4	
+1	+	x	+	(−2) 4	= 0
+1	+	x	+	− 8	= 0
		x	+	-7	= 0
		x	=	**+ 7**	

Example 12.2:

Calculate the oxidation number of Cr is $K_2\,Cr_2\,O_7$.

Solution: K_2 $\quad$ Cr_2 $\quad$ O_7

(+1) 2 + (x) 2 + (−2) 7

$$+2 \quad + \quad 2x \quad + \quad -14 = 0$$
$$+ \quad 2x \quad + \quad -12 = 0$$
$$2x = 12$$
$$x = 12/2 = 6$$

Thus oxidation number of Cr in $K_2Cr_2O_7$ is +6

Example 12.3:

What is the oxidation state of S in SO_4^{-2} ion.

Solution:

In SO_4^{-2} ion oxidation state of O is –2. If x is the oxidation state of S then.

$$x + 4(-2) = -2$$
$$x - 8 = -2$$
$$x = 6$$

SELF-CHECK EXERCISE 12.1

Identify the compound in which oxidation number of Fe is +3

FeO, Fe_3O_4, Fe_2O_3.

Solution:

FeO:

$Oxidation\ number\ of\ O = -2$

$oxidation\ number\ of\ Fe = x$

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

Fe_3O_4:

$Oxidation\ number\ of\ O = -2$

$oxidation\ number\ of\ Fe_3 = 3x$

$$3x + 4(-2) = 0$$
$$3x - 8 = 0$$
$$3x = +8$$
$$x = +\frac{8}{3}$$

Fe_2O_3:

$Oxidation\ number\ of\ O = -2$

$oxidation\ number\ of\ Fe_2 = 2x$

$$2x + 3(-2) = 0$$
$$2x - 6 = 0$$
$$2x = +6$$
$$x = +3$$

$Thus,\ in\ Fe_2O_3,\ the\ oxidation\ state\ of\ Fe\ is\ +3$

Q6. How oxidation-reduction can be explain in terms of change in oxidation number?

Ans: Oxidation-Reduction in Terms of Change in Oxidation Number:

We can also define oxidation and reduction in terms of change in oxidation number.

Oxidation is an increase in oxidation number (a loss of electrons).

Reduction is a decrease in oxidation number (a gain of electrons).

Explanation:

Consider the following reaction

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

Assign oxidation number to all the atoms involved in this reaction and write it over their symbols

$$2\overset{0}{Na}(s) + \overset{0}{Cl_2}(g) \longrightarrow 2\overset{+1}{Na}\overset{-1}{Cl}(s)$$

Notice that the oxidation number of Na is zero because it is in the elemental form. In this reaction Na undergoes a change in oxidation number from zero in Na to +1 in NaCl. Thus this change can be accounted for by a loss of one electron per Na atom. This is oxidation.

$$2\overset{0}{Na}(s) \longrightarrow 2\overset{+1}{Na} + 2e^-$$

On the other hand, each Cl atom in Cl_2 molecule changes its oxidation number from zero in Cl_2 to –1 in NaCl. In this change gain of one electron per Cl atom occurs. This is reduction

$$\overset{0}{Cl_2} + 2e \longrightarrow 2Cl^{-1}$$

Conclusion:

Thus we can also define oxidation and reduction in terms of change in oxidation number

Oxidation is an increase in oxidation number (a loss of electrons).

Reduction is a decrease in oxidation number (a gain of electrons).

Example 12.4:

Identify the elements undergoing oxidation or reduction in terms of change in oxidation number in the following reactions which takes place in the combustion of natural gas.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O\ (l)$$

Solution: Assign oxidation number to all the atoms involved in this reaction.

$$\overset{-4}{C}\overset{+1}{H_4}(g) + 2\overset{0}{O_2}(g) \longrightarrow \overset{+4}{C}\overset{-2}{O_2}(g) + 2\overset{+1}{H_2}\overset{-2}{O}(l)$$

The C changes its oxidation number from –4 in CH_4 to +4 in CO_2. This is 8 electrons loss. This means C undergoes an increase in oxidation number. On the other hand, O changes its oxidation number from zero in O_2 to –2 in H_2O and CO_2. Each oxygen atom gains two electrons and therefore it is reduced.

We can say that.

i. C is oxidized because there has been an increase in its oxidation number.

ii. O is reduced because there has been a decrease in its oxidation number.

Example 12.5:

Identify the substance oxidized and the substance reduced in the dry cell.

Solution: Following reaction can take place in dry cell.

At Anode. $Zn \longrightarrow Zn^{+2} + 2e^-$

At Cathode. $2NH_4^+ + 2MnO_2 + 2e^- \longrightarrow Mn_2O_3 + 2NH_3 + H_2O$

Anode reaction indicates that Zn metal lose 2 electrons, therefore it undergoes oxidation. For determining the substance reduced, assign oxidation number to all the atoms involved in the reaction that occurs at cathode

$$\overset{-3(+1)4}{2[NH_4]^+} \quad \overset{+4}{2MnO_2} \quad 2e \quad \longrightarrow \quad \overset{+3}{Mn_2O_3} \quad 2NH_3 \quad H_2O$$

In this reaction Mn changes its oxidation number from +4 in MnO_2 to +3 in Mn_2O_3

This means that each Mn gains one electron therefore it undergoes reduction. Thus the substance reduced is Mn

SELF-CHECK EXERCISE 12.2

Use the oxidation number change method to identify the atoms undergoing oxidation or reduction in the following redox reactions.

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$
$$4Fe(s) + 3O_2(g) \longrightarrow 2Fe_2O_3(s)$$
$$C_6H_{12}O_6(s) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l)$$

Solution: $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$

Assigning oxidation number:

$$\overset{0}{N_2(g)} + \overset{0}{3H_2(g)} \longrightarrow \overset{-3\ 3(+1)}{2NH_3(g)}$$

N_2 changes its oxidation number from 0 to -3 in NH_3. This is due to gain of 3 electrons. Thus it is reduction.

H_2 changes its oxidation number from 0 to +1 in NH_3. This is due to loss of one electron. Thus it is oxidation.

$4Fe(s) + 3O_2(g) \longrightarrow 2Fe_2O_3(s)$

Assigning oxidation number:

$$\overset{0}{4Fe(s)} + \overset{0}{3O_2(g)} \longrightarrow \overset{2(+3)\ 3(-2)}{2Fe_2O_3(s)}$$

The Fe atom changes its oxidation number from zero to +3 in Fe_2O_3. It is because of the loss of 3 electrons So this is oxidation.

The O atom changes its oxidation number from zero to -2 in Fe_2O_3. It is because of the gain of 2. Thus it is reduction.

$C_6H_{12}O_6(s) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l)$

Assigning oxidation number:

$$\overset{0\ 12(+1)\ 6(-2)}{C_6H_{12}O_6(s)} + \overset{0}{6O_2(g)} \longrightarrow \overset{+4\ 2(-2)}{6CO_2(g)} + \overset{2(+1)\ -2}{6H_2O(l)}$$

The C atom changes its oxidation number from 0 to +4 in CO_2. It is because of the loss of 4 electrons. Thus this is oxidation.

The O atom changes its oxidation number from 0 to -2 in CO_2 and H_2O. It is because of the gain of 2 electrons. Hence it is reduction.

Q7. How can we balance redox equations by oxidation number change method?

Ans: Balancing Redox Equations by Oxidation Number Change Method:

It is based on the principle that in any redox reaction, the total number of electrons lost by one element must be equal to the total number of electrons gained by another element. This method can be understood by the following example.

Example:

Balance the following equation by oxidation number method.

$$P + HNO_3 + H_2O \longrightarrow H_3PO_4 + NO$$

Step1: Assign oxidation number to all the atoms involved in the equation.

0   +1 +5 (-2) 3   (+1)2 -2   3(+1) +5 (-2)4   2 -2

$P +$ H N O_3 $+ H_2O \longrightarrow H_3$ P O_4 + N O

Step2: Identify the elements undergoing a change in oxidation number.

The P goes from zero to +5 oxidation state in H_3PO_4. This is a 5 electron change. N in HNO_3 goes from +5 to +2 oxidation state in NO. This is 3 electron change.

Step3: Draw a bridge between the same atoms whose oxidation number have changed, Indicate this change by the number of electrons gained or lost by each element.

$+3e^-$

$$\overset{0}{P} + \overset{+5}{HNO_3} \longrightarrow \overset{+5}{H_3PO_4} + \overset{-2}{NO}$$

$-5e^-$

Step4: Equalize the number of electrons lost and gained by multiplying the two numbers, by a small whole number which produces a common number. Use these multipliers as coefficients of the respective substance.

To balance a $3e^-$ gain against a $5e^-$ loss, we need to multiply $3e^-$ gain by 5 and $5e^-$ loss by 3. This will equalize the number of electrons gained and lost.

$+3e^- \times 5 = +15e^-$

$$\overset{0}{P} + \overset{+5}{HNO_3} \longrightarrow \overset{+5}{H_3PO_4} + \overset{-2}{NO}$$

$-5e^- \times 3 = -15e^-$

Multiply the coefficients of P and that of H_3PO_4 by 3. Whereas multiply coefficients of HNO_3 and NO by 5.

$$3P + 5HNO_3 + H_2O \longrightarrow 3H_3PO_4 + 5NO$$

Now the coefficient of H_3PO_4 and NO should not be changed hereafter it.

Step5: Balance the rest of the equation by inspection method. Balance the atoms other than oxygen and hydrogen first, then oxygen atoms and finally hydrogen atoms.

To balance oxygen atoms multiply coefficient of H_2O by 2.

$$3P + 5HNO_3 + 2H_2O \longrightarrow 3H_3PO_4 + 5NO$$

Inspect the equation, it is balanced.

SELF-CHECK EXERCISE 12.3

Using the oxidation number method balance the following equation.

$$HNO_3 + H_2S \longrightarrow NO + S + H_2O$$

Solution:

Step 1: Assign oxidation number to all the atoms involved in the equation.

$$\overset{+1}{H}\overset{+5}{N}\overset{3(-2)}{O_3} + \overset{2(+1)\,-2}{H_2S} \longrightarrow \overset{+2\,-2}{NO} + \overset{0}{S} + \overset{2(+1)\,-2}{H_2O}$$

Step2: Identify the elements undergoing a change in oxidation number

N atom changes its oxidation state from +5 to +2 in NO. It is because of 3 electron change.

The S atom changes its oxidation state from -2 to zero in S. It is because of the 2 electron change.

Step3: Draw a bridge between the same atoms undergoing a change In oxidation state. Indicate this change by the number of electrons gained or lost by each element.

$-2e^-$ (loss of $2e^-$: Oxidation)

$$\overset{+1}{H}\overset{+5}{N}\overset{3(-2)}{O_3} + \overset{2(+1)\,-2}{H_2S} \longrightarrow \overset{+2\,-2}{NO} + \overset{0}{S} + \overset{2(+1)\,-2}{H_2O}$$

$+3e^-$ (gain of $3e^-$: Reduction

Step4: Balance the loss and gain of electrons by multiplying with small suitable whole numbers. Use these multipliers as coefficient for respective substances.

To balance $3e^-$ gain against $2e^-$ loss multiply $3e^-$ gain by 2 and $2e^-$ loss by 3. This will equalize the number of electron gained and lost.

$-2e^- \times \boxed{3} = -6e^-$

$$\overset{+1}{H}\overset{+5}{N}\overset{3(-2)}{O_3} + \overset{2(+1)\,-2}{H_2S} \longrightarrow \overset{+2\,-2}{NO} + \overset{0}{S} + \overset{2(+1)\,-2}{H_2O}$$

$+3e^- \times \boxed{2} = +6e^-$

Multiply the coefficients of H_2S and that of S by 3. Whereas multiply coefficients of HNO_3 and NO by 2.

$$2HNO_3 + 3H_2S \longrightarrow 2NO + 3S + H_2O$$

These coefficients should not be changed after this.

Step5: Balance the rest of the equation by Inspection method. Balance the atoms other than oxygen and hydrogen first, then oxygen atoms and finally hydrogen atoms.

To balance oxygen atoms multiply coefficient of H_2O by 4

$$2HNO_3 + 3H_2S \longrightarrow 2NO + 3S + 4H_2O$$

This is the balanced equation.

Q8. How a redox reaction can be separated into two halves explain your answer with the help of example.

Ans: Breaking a Redox Reaction into Oxidation and Reduction Reactions:

A redox reaction in aqueous reaction can be separated into two half-reactions. One, half reaction represents oxidation and the other represents reduction.

A **half-reaction** is the part of an overall redox reaction that represents, separately, either an oxidation or a reduction.

Explanation:

Consider the following reaction which is used to analyze iron ore for its iron content.

$$MnO_4^{-1} + Fe^{+2} \xrightarrow{Acid} Fe^{+3} + Mn^{+2}$$

Step1: Write oxidation number of all the atoms involved in this reaction.

$$\left[\overset{+7}{Mn}\ \overset{(-2)4}{O_4}\right]^{-1} + Fe^{+2} \xrightarrow{Acid} Fe^{+3} + Mn^{+2}$$

Step2: Identify and write equation for the half-reactions.

Notice that the Mn goes from +7 oxidation state in MnO_4^{-1} to +2. The reduction half-reaction must involve species containing Mn.

$MnO_4^{-1} \longrightarrow Mn^{+2}$ (Reduction half-reaction)

Iron goes from+2 to +3 oxidation state. The oxidation half-reaction must involve Fe^{+2} and Fe^{+3}

$Fe^{+2} \longrightarrow Fe^{+3}$ (Oxidation half-reaction)

Example 12.7:

Split the following reaction into half-reactions.

$$SO_2 + MnO_4^{-1} \longrightarrow SO_4^{-2} + Mn^{+2}$$

Solution:

decrease in oxidation number (reduction)

$$\overset{+4}{S}\overset{(-2)2}{O_2} + \left[\overset{+7}{Mn}\ \overset{(-2)4}{O_4}\right]^{-1} \longrightarrow \left[\overset{+6}{S}\overset{(-2)4}{O_4}\right]^{-2} + Mn^{+2}$$

Increase in oxidation number (oxidation)

Reduction half reaction: $MnO_4^{-1} \longrightarrow Mn^{+2}$

Oxidation half reaction: $SO_2 \longrightarrow SO_4^{-2}$

SELF-CHECK EXERCISE 12.4

Split the following reactions into two half reactions:

(i) $Mn^{+2} + PbO_2 \longrightarrow MnO_4^{-1} + Pb^{+2}$

(ii) $Br_2 + SO_2 \longrightarrow SO_4^{-2} + Br^{-1}$

(iii) $NO + MnO_4 \longrightarrow NO_3 + MnO_2$

Solution:

(i) $Mn^{+2} + PbO_2 \longrightarrow M_nO_4^{-1} + Pb^{+2}$

Decrease in oxidation number (***Reduction***)

$$\overset{+2}{Mn} + \overset{+4}{Pb}\,\overset{2(-2)}{O_2} \rightarrow \left[\overset{+7}{Mn}\,\overset{4(-2)}{O_4}\right]^{-1} + \overset{+2}{Pb}$$

Increase in oxidation number (***Oxidation***)

Reduction half reaction:

$$PbO_2 \longrightarrow Pb^{+2}$$

Oxidation half reaction:

$$Mn^{+2} \longrightarrow M_nO_4^{-1}$$

(ii) $Br_2 + SO_2 \longrightarrow SO_4^{-2} + Br^{-1}$

Solution:

Decrease in oxidation number (***Reduction***)

$$\overset{0}{Br_2} + \overset{+4}{S}\,\overset{2(-2)}{O_2} \rightarrow \left[\overset{+6}{S}\,\overset{4(-2)}{O_4}\right]^{-2} + \overset{-1}{Br}$$

Increase in oxidation number (***Oxidation***)

Reduction half reaction:

$$Br_2 \longrightarrow Br^{-1}$$

Oxidation half reaction:

$$SO_2 \longrightarrow SO_4^{-2}$$

(iii) $NO + MnO_4^{-1} \longrightarrow NO_3^{-1} + MnO_2$

Solution:

Decrease in oxidation number (***Reduction***)

$$\overset{+2}{N}\overset{-2}{O} + \left[\overset{+7}{Mn}\,\overset{4(-2)}{O_4}\right]^{-1} \rightarrow \left[\overset{+5}{N}\,\overset{3(-2)}{O_3}\right]^{-1} + \overset{+4}{Mn}\overset{2(-2)}{O_2}$$

Increase in oxidation number (***Oxidation***)

Reduction half reaction:

$$MnO_4^{-1} \longrightarrow MnO_2$$

Oxidation half reaction:

$$NO \longrightarrow NO_3^{-1}$$

Q9. How can we balance a redox reaction with the help of half reaction method?

Ans: Half Reaction Method to Balance a Redox Reaction:

A powerful technique for balancing redox reactions involves dividing these reactions into separate oxidation and reduction half reactions.

We then balance the half-reaction, one at a time. And combine them so that electrons are neither created nor destroyed in the reaction.

The steps involved in this method can be understood by considering the following reactions used to determine the amount of the tri-iodide ion (I_3^-) in a solution by titration.

$$I_3^- + S_2O_3^{-2} \longrightarrow I^- + S_4O_6^{-2}$$

Step1: Break the reaction into oxidation and reduction half-reactions.

Increase in oxidation number (oxidation)

$$\left[\overset{-1/3\times 3}{I_3}\right]^- + \left[\overset{2(+2)}{S_2}\ \overset{(-2)3}{O_3}\right]^{-2} \longrightarrow I^- + \left[\overset{4(5/2)}{S_4}\ \overset{(-2)6}{O_6}\right]^{-2}$$

decrease in oxidation number (reduction)

Reduction half-reaction: $I_3^- \longrightarrow I^-$

Oxidation half-reaction: $S_2O_3^{-2} \longrightarrow S_4O_6^{-2}$

Step2: Balance these half-reactions one at a time. It doesn't matter which half-reaction we balance first. Balance each half reaction in terms of both charge and mass. Let us start with the reduction half reaction.

$$I_3^- \longrightarrow I^-$$

There are 3 I atoms on the left side and one on the right side. So we will multiply co-efficient of I^- by 3. This will balance I atoms on both the sides.

$$I_3^- \longrightarrow 3I^-$$

Now balances charges. The right side has 3 mono negative charges corresponding to -3 charges. But the left side has only one uni-negative charge (-1). Thus left side needs $2e^-$

$$I_3^- + 2e^- \longrightarrow 3I^- \qquad \text{............}\mathbf{(1)}$$

Now balance oxidation half reaction.

$$S_2O_3^{-2} \longrightarrow S_4O_6^{-2}$$

To balance S atoms on both the sides, multiply co-efficient of $S_2O_3^{-2}$ by 2. This will also balance O atoms.

$$2S_2O_3^{-2} \longrightarrow S_4O_6^{-2}$$

Now balance charges. The left side has two di negative charges corresponding to $2\times(-2) = -4$. The right side has -2 charges. Thus the right side needs $2e^-$.

$$2S_2O_3^{-2} \longrightarrow S_4O_6^{-2} + 2e^- \qquad \text{......................}\mathbf{(2)}$$

Notice that both the half reaction are balanced in terms of both mass and charge.

Step3: Combine these half reactions so that electrons are neither created nor destroyed. Reduction half reaction uses up $2e^-$ and oxidation half reaction produces $2e^-$, we can therefore obtain a balanced equation by simply adding equation (1) and equation (2)

$$I_3^- + 2e^- \longrightarrow 3I^-$$

$$2S_2O_3^{-2} \longrightarrow S_4O_6^{-2} + 2e^-$$

$$I_3^- + 2e^- + 2S_2O_3^{-2} \longrightarrow 3I^- + S_4O_6^{-2} + 2e^-$$

Step4: Cancel Duplication of Species on both the sides.

Duplications are 2e Strike these out from both the sides.

$$I_3^- + 2S_2O_3^{-2} \longrightarrow 3I^- + S_4O_6^{-2}$$

Check, the overall reaction is balanced in terms of both charge and mass.

SELF-CHECK EXERCISE 12.5

Use the half reaction method to balance the following redox reactions.

(i)	$Co + Cr^{+3}$	$\longrightarrow$	Co^{+2}	+	Cr^{+2}
(ii)	$Au^{+3} + H_2$	$\longrightarrow$	H^+	+	Au
(iii)	$Cl_2 + I^-$	$\longrightarrow$	I_2	+	Cl^-

Solution: **(i)** $Co + Cr^{+3} \longrightarrow Co^{+2} + Cr^{+2}$

Step 1: Break the reaction into oxidation and reduction half-reactions.

Decrease in oxidation number (***Reduction***)

$$\overset{0}{Co} + \overset{+3}{Cr} \rightarrow \overset{+2}{Co} + \overset{+2}{Cr}$$

Increase in oxidation number (***Oxidation***)

Reduction half-reaction:

$$Cr^{+3} \longrightarrow Cr^{+2}$$

Oxidation half-reaction:

$$Co \longrightarrow Co^{+2}$$

Step 2: Balance these half-reactions one at a time. It doesn't matter which half-reaction we balance first. Balance each half reaction in terms of both charge and mass.

First consider the reduction half-reaction.

$$Cr^{+3} \longrightarrow Cr^{+2}$$

Balances Charge:

The left side has one tri-positive charge corresponding to +3. The right side has one di-positive charge corresponding to +2 Thus, left side needs 1e⁻

$$Cr^{+3} + 1e^- \longrightarrow Cr^{+2} \quad \text{.......................(i)}$$

Oxidation half-reaction:

$$Co \longrightarrow Co^{+2}$$

Balances Charge:

Left side has no charge. Right side has one di-positive charge correspond +2 Thus the R.H.S. needs 2e to balance di-positive charge.

$$Co \longrightarrow Co^{+2} + 2e^- \quad \text{.......................(ii)}$$

Both the half reactions are now balanced.

Step 3: Combine these half reactions so that electrons are neither created nor destroyed.

The reduction half-reaction uses up $1e^-$ while the oxidation half-reaction produces $2e^-$ we can obtain balanced equation by multiply Eq (i) by 2 and then add it to Eq (ii).

$$2Cr^{+3} \quad + \quad 2e^- \longrightarrow \quad 2Cr^{+2}$$
$$\underline{Co \quad \longrightarrow \quad Co^{+2} \quad + \quad 2e^-}$$
$$Co + 2Cr^{+3} + 2e^- \longrightarrow Co^{+2} + 2Cr^{+2} + 2e^-$$

Step 4: Cancel Duplication of Species on both the sides.

Duplications are $2e^-$. Strike these out from both the sides.

$$Co + 2Cr^{+3} \longrightarrow Co^{+2} + 2Cr^{+2}$$

Check, the overall reaction is balanced in terms of both charge and mass.

(ii) $Au^{+3} + H_2 \longrightarrow H^+ + Au$

Step 1: Break the reaction into oxidation and reduction half-reactions.

Decrease in oxidation number (***Reduction***)

$$\overset{+3}{Au} \quad + \quad \overset{0}{H_2} \quad \rightarrow \quad \overset{+1}{H} \quad + \overset{0}{Au}$$

Increase in oxidation number (***Oxidation***)

Reduction half-reaction: $Au^{+3} \longrightarrow Au$

Oxidation half-reaction: $H_2 \longrightarrow H^{+1}$

Step 2: Balance these half-reactions one at a time. It doesn't matter which half-reaction we balance first. Balance each half reaction in terms of both charge and mass.

Reduction half-reaction:

$$Au^{+3} \longrightarrow Au$$

Balance Mass: Mass is already balanced.

Balances Charge:

The left side has one tri-positive charge corresponding to +3. The right side has no charge. Thus, left side needs $3e^-$.

$$Au^{+3} + 3e^- \longrightarrow Au \quad \text{..............................(i)}$$

Oxidation half reaction:

$$H_2 \longrightarrow H^{+1}$$

Balance Mass: There are two hydrogen atoms on the left side and one on the right side so we will multiply 2 by H^{+1}.

$$H_2 \longrightarrow 2H^{+1}$$

Balance Charge:

Left side has no charge. Right side has two mono-positive charges correspond to +2. Thus the right side needs 2e

$$H_2 \longrightarrow 2H^{+1} + 2e^- \quad \text{....(ii)}$$

Step 3: Combine these half reactions so that electrons are neither created nor destroyed.

Reduction half-reaction uses up 3e⁻ and oxidation half-reaction produces 2e⁻. We can therefore obtain a balance equation by multiply Eq. (i) by (ii), Eq. (ii) by (iii) and then add the resulting equations.

$$2Au^{+3} + 6e^- \longrightarrow 2Au$$
$$3H_2 \longrightarrow 6H^{+1} + 6e^-$$

$$\overline{2Au^{+3} + 3H_2 + 6e^- \longrightarrow 6H^{+1} + 2Au + 6e^-}$$

Step 4: Cancel Duplication of Species on both the sides.

Duplications are $6e^-$. Strike these out from both the sides.

$$2Au^{+3} + 3H_2 \longrightarrow 6H^{+1} + 2Au$$

(iii) $Cl_2 + I^{-1} \longrightarrow I_2 + Cl^{-1}$

Step 1: Split the reaction into oxidation and reduction half-reactions.

Decrease in oxidation number (***Reduction***)

$$\overset{0}{Cl_2} + \overset{-1}{I} \rightarrow \overset{0}{I_2} + \overset{-1}{Cl}$$

Increase in oxidation number (***Oxidation***)

Reduction half-reaction:

$$Cl_2 \longrightarrow Cl^{-1}$$

Oxidation half-reaction:

$$I^{-1} \longrightarrow I_2$$

Step 2: Balance these half-reactions one at a time. It doesn't matter which half-reaction we balance first. Balance each half reaction in terms of both charge and mass.

Reduction half-reaction:

$$Cl_2 \longrightarrow Cl^{-1}$$

Balance Mass: There are two Cl atom on the left side and one on the right side so we will multiply Cl^{-1} by 2.

$$Cl_2 \longrightarrow 2Cl^{-1}$$

Balances Charge:

Left side has no charge. Right side has two uni-negative charges correspond to -2. Thus, left side needs $2e^-$.

$$Cl_2 + 2e^- \longrightarrow 2Cl^{-1} \text{ ....(i)}$$

Oxidation half reaction:

$$I^{-1} \longrightarrow I_2$$

Balance Mass: There are 2 atoms of I on the right side and 1 on the left side so multiply I^{-1} by 2.

$$2I^{-1} \longrightarrow I_2$$

Balance Charge:

Left side has two uni-negative charges correspond to -2. Right side has no charge. Thus right side needs $2e^-$.

$$2I^{-1} \longrightarrow I_2 + 2e^- \text{ ....(ii)}$$

Step 3: Combine these half reactions so that electrons are neither created nor destroyed.

Reduction half-reaction uses up 2e and oxidation half-reaction produces 2e we can obtain a balance equation by simply adding Eq. (i) and Eq. (ii).

$$Cl_2 + 2e^- \longrightarrow 2Cl^{-1}$$
$$2I^{-1} \longrightarrow I_2 + 2e$$

$$Cl_2 + 2I^{-1} + 2e^- \longrightarrow I_2 + 2Cl^{-1} + 2e^-$$

Step 4: Cancel Duplication of Species on both the sides.

Duplications are $6e^-$. Strike these out from both the sides.

$$Cl_2 + 2I^{-1} \longrightarrow I_2 + 2Cl^{-1}$$

Half Reaction Method to Balance a Redox Reaction in Acid Solution:

We can understand this method by the following example.

Example 12.8:

Balance the following equation by half reaction method.

$$Cr_2O_7^{-2} + H_2SO_3 \longrightarrow Cr^{+3} + HSO_4^-$$

Solution:

Step1: Split the reaction into two half reactions.

Reduction half reaction:

$$Cr_2O_7^{-2} \longrightarrow Cr^{+3}$$

Oxidation half reaction:

$$H_2SO_3 \longrightarrow HSO_4^-$$

Step2:

Balance each half reaction. First consider reduction half reaction. Two Cr atoms on the left require 2 before Cr^{+3}

$$Cr_2O_7^{-2} \longrightarrow 2Cr^{+3}$$

There are seven O atoms on the left and none on the right. So we will add $7H_2O$ on the right side.

$$Cr_2O_7^{-2} \longrightarrow 2Cr^{+3} + 7H_2O$$

There are 14 H atoms on the right and none on the left, so we will add 14 H^+ on the left side.

$$Cr_2O_7^{-2} + 14H^+ \longrightarrow 2Cr^{+3} + 7H_2O$$

Now balance charges. The left side has one di-negative and 14 mono-positive, corresponding to $-2 + 14 = +12$. The right side has two tri-positive corresponding to $+3 \times 2 = +6$. Thus left side needs $6e^-$.

$$Cr_2O_7^{-2} + 14H^+ + 6e^- \longrightarrow Cr^{+3} + 7H_2O \quad \text{---(1)}$$

In the other half reaction (Oxidation half reaction), S atoms are already balanced.

$$H_2SO_3 \longrightarrow HSO_4^-$$

Balance O – atoms. As there are three O – atoms on the left and four on the right, we will add one H_2O to the left.

$$H_2SO_3 + H_2O \longrightarrow HSO_4^-$$

There are four H-atoms on the left and one on the right. We will add $3H^+$ to the right.

$$H_2SO_3 + H_2O \longrightarrow HSO_4^- + 3H^+$$

For charge, the left side is neutral but the right side has a net charge of

$(-1) + (+3) = +2$

Thus we will add 2e to the right side.

$$H_2SO_3 + H_2O \longrightarrow HSO_4 + 3H^+ + 2e \quad \text{-----(2)}$$

Step3:

Equalize the number of electrons transferred in the two half reactions and add half reactions. Reduction half reaction uses up 6e and oxidation half reaction produce 2e Therefore multiplying equation (1) by one and equation (2) by three and adding two equations we get.

$$Cr_2O_7^{-2} + 14H^+ + 6e \longrightarrow 2Cr^{+3} + 7H_2O$$

$$3H_2SO_3 + 3H_2O \longrightarrow 3HSO_4 + 9H^+ + 6e^-$$

$$Cr_2O_7^{-2} + 14H^+ + 3H_2SO_3 + 3H_2O + 6e^- \longrightarrow 2Cr^{+3} + 7H_2O + 3HSO_4 + 9H^+ + 6e^-$$

Step4:

Cancel the duplication. Duplications are $6e^-$, $3H_2O$ and $9H^+$ Strike these out from both sides.

$$Cr_2O_7 + 5H^+ + 3H_2SO_3 \longrightarrow 2Cr^{+3} + 4H_2O + 3HSO_4$$

Example 12.9:

Use the half reaction method to balance the following reaction that takes place in acidic medium.

$$\mathbf{MnO_4^- + H_2O_2 \longrightarrow Mn^{+2} + O_2}$$

Solution:

Step1: Split the reaction into two half-reactions.

Reduction half reaction:

$$MnO_4^- \longrightarrow Mn^{+2}$$

Oxidation half reaction:

$$H_2O_2 \longrightarrow O_2$$

Step2: Balance each half reaction. First consider reduction half reaction.

$$MnO_4^- \longrightarrow Mn^{+2}$$

Mn atoms on both the sides are already balanced. There are four O atoms on the left side and non on the right. so we will add four $4H_2O$ on the right side.

$$MnO_4^- \longrightarrow Mn^{+2} + 4H_2O$$

There are 8 H atoms on the right side and non on the left so. we will add $8H^+$ on the left side.

$$MnO_4 + 8H^+ \longrightarrow Mn^{+2} + 4H_2O$$

Step3:

Now balance charges. The left side has one uni-negative and 8 mono-positive charges corresponding to $-1+8 = +7$ The right side has one di-positive corresponding to +2. Thus left side needs $5e^-$.

$$MnO_4 + 8H^+ + 5e^- \longrightarrow Mn^{+2} + 4H_2O \qquad (1)$$

Now consider oxidation half-reaction.

$$H_2O_2 \longrightarrow O_2$$

O atoms on both sides are equal. There are 2H atoms on the left side and non on the right side. So we will add $2H^+$ on the right side

$$H_2O_2 \longrightarrow O_2 + 2H^+$$

Now balance charges. The left side is neutral whereas the right side has two mono positive charge corresponding to $+1\times2 = +2$. Thus the right side needs 2e.

$$H_2O_2 \longrightarrow O_2 + 2H^+ + 2e \qquad (2)$$

Step4:

Reduction half reaction uses up 5e and oxidation half reaction produces 2e Therefore, multiply equation (1) by 2 and equation (2) by 5 and then add these equations.

$$2MnO_4^- + 16H^+ + 10e \longrightarrow 2Mn^{+2} + 8H_2O$$

$$5H_2O_2 \longrightarrow 5O_2 + 10H^+ + 10e$$

$$2MnO_4^- + 16H^+ + 10e + 5H_2O_2 \longrightarrow 2Mn^{+2} + 8H_2O + 5O_2 + 10H^+ + 10e$$

Step5:

Strike out duplications from both the sides.

$$2MnO_4^- + 6H^+ + 5H_2O_2 \longrightarrow Mn^{+2} + 8H_2O + 5O_2$$

SELF-CHECK EXERCISE 12.6

1. Balance each of the following half reactions that take place in acidic medium.

a) $NO^{3-} \longrightarrow NO_2$

b) $IO^{3-} \longrightarrow I_2$

c) $ClO^{4-} \longrightarrow ClO^{3-}$

2. Balance the following reactions by half–reaction method, which take place in acidic medium.

a) $NO_3^- + Br^- \longrightarrow NO + Br_2$

b) $Ce^{+4} + H_3AsO_3 \longrightarrow Ce^{+3} + H_3AsO_4$

c) $S_2O_8\text{-}2 + Cr\text{+}3 \longrightarrow SO_4^{-2} + Cr_2O_7^{-2}$

d) $MnO^{4-} + C_2O_4^{-2} \longrightarrow Mn^{+2} + CO_2$

Solution:

1. Balance each of the following half reactions that take place in acidic medium.

a) $NO_3^- \longrightarrow NO_2$

Step 1:

N atoms on the both side are already balanced. There are three O atoms on the left side and two on the right side so we will add one H_2O on the right side.

$$NO_3^- \longrightarrow NO_2 + H_2O$$

There are 2 H atoms on the right side and none on the left side so we will add $2H^+$ on the left side.

$$NO_3^- + 2H^+ \longrightarrow NO_2 + H_2O$$

Step 2:

Now balance charges right side has no charge and the left side has uni-negative and mono-positive charges corresponding to -1+2(+1) = +1. Thus left side needs $1e^-$

$$NO_3^- + 2H^+ + 1e^- \longrightarrow NO_2 + H_2O$$

This is the balanced reduction half reaction.

b) IO_3^- $\longrightarrow$ I_2

Reduction half reaction:

Step 1:

There is one I atom on the left side and 2 on the right side so we will balance this by multiplying IO_3^- with 2

$$2IO_3 \longrightarrow I_2$$

As there are 6 oxygen atoms on the left side and none on the right side so we will add $6H_2O$ on the right side

$$2IO_3 \longrightarrow I_2 + 6H_2O$$

There are 12 H atom on the right side and none on the left side so we will add $12H^+$ on the left side.

$$2IO_3^- + 12H^+ \longrightarrow I_2 + 6H_2O$$

Step 2:

Now balance charges there is no charge on the right side and there are 12 mono-positive and 1 uni-negative charge on the left side corresponding to 2(-1)+12(+1) = +10. Thus, left side needs $10e^-$

$$2IO_3^- + 12H^+ + 10e^- \longrightarrow I_2 + 6H_2O$$

This is balanced reduction half reaction.

c) ClO_4^- $\longrightarrow$ ClO_3^-

Step 1:

Cl atoms on the both sides are already balanced.

$$ClO_4^- \longrightarrow ClO_3^-$$

There are 4 oxygen atoms on the left side and three on the right side so we will add 1 H_2O on the right side.

$$ClO_4^- \longrightarrow ClO_3^- + H_2O$$

There are 2 hydrogen atoms on the right side and none on the left side so we will add $2H^+$ on left side.

$$ClO_4^- + 2H^+ \longrightarrow ClO_3^- + H_2O$$

Step 2:

$$ClO_4^- + 2H^+ \longrightarrow ClO_3^- + H_2O$$

Now balance charges there is 1 uni-negative charge on the right side corresponding to -1 and there are 2 mono-positive and 1 uni-negative charges on the left side corresponding to -1+2(+1) = +1. Thus, left side needs $2e^-$.

$$ClO_4^- + 2H^+ + 2e^- \longrightarrow ClO_3 + H_2O$$

This is the balanced reduction half reaction.

2. Balance the following reactions by half–reaction method, which take place in acidic medium.

a) NO_3^- + Br^- $\longrightarrow$ NO + Br_2

Step 1: Split the reaction into two half-reactions.

Reduction half-reaction: $NO_3^- \longrightarrow NO$

Oxidation half-reaction: $Br^- \longrightarrow Br_2$

Step 2: Balance each half-reaction separately.

First consider reduction half-reaction nitrogen is already balanced there are 3 oxygen atoms on the left side and 1 on the right side so add $2H_2O$ on the right side.

$$NO_3^- \longrightarrow NO + 2H_2O$$

There are 4 hydrogen atoms on the right side and none on the left side so we will add $4H^+$ on the left side.

$$NO_3^- + 4H^+ \longrightarrow NO + 2H_2O$$

Balance Charge:

Now balance charges there are 4 mono-positive charge and 1 uni-negative charge on the left side corresponding to -1+4(+1) = +3 and no charge is present on the right side. Thus, left side needs $3e^-$.

$$NO_3^- + 4H^+ + 3e^- \longrightarrow NO + 2H_2O \quad \text{.....................(i)}$$

Oxidation half reaction:

Balance Mass:

$$Br^- \longrightarrow Br_2$$

There are 2Br atoms on the right side and 1 on the left side to balance this multiply Br^- with 2

$$2Br^- \longrightarrow Br_2$$

Balance Charge:

Now balance charges there is 1 uni-negative charge, on the left side corresponding to 2(-1) = -2 and none on the right side. Thus, right side needs $2e^-$

$$2Br^- \longrightarrow Br_2 + 2e^- \quad \text{...........................(ii)}$$

Step 3:

Balance the loss and gain of electrons in both half reactions and add them. The reduction half-reaction has $3e^-$ while the oxidation half-reaction has $2e^-$. Thus, multiply Eq. (i) by (ii), Eq. (ii) by (iii) and then add the resulting equations.

$$2NO_3^- + 8H^+ + 6e^- \longrightarrow 2NO + 4H_2O$$
$$6Br^- \longrightarrow 3Br_2 + 6e^-$$

$$2NO_3^- + 8H^+ + 6e^- + 6Br^- \longrightarrow 2NO + 4H_2O + 3Br_2 + 6e^-$$

Step 4:

Cancel Duplication of Species on both the sides.

Duplications are $6e^-$. Strike out from both the sides.

$$2NO_3^- + 8H^+ + 6Br^- \longrightarrow 2NO + 4H_2O + 3Br_2$$

b) $Ce^{+4} + H_3AsO_3 \longrightarrow Ce^{+3} + H_3AsO_4$

Step 1: Split the reaction into two half-reactions.

Reduction half-reaction: $Ce^{+4} \longrightarrow Ce^{+3}$

Oxidation half-reaction: $H_3AsO_3 \longrightarrow H_3AsO_4$

Step 2: Balance each half-reaction separately.

Reduction half-reaction:

Ce atoms are already balanced.

$$Ce^{+4} \longrightarrow Ce^{+3}$$

Balance Charge:

Now balance charges there are 4 positive charges on the left side corresponding to +4. There are 3 positive charges on the right side corresponding to +3. Thus, add $1e^-$ on L.H.S.

$$Ce^{+4} + 1e^- \longrightarrow Ce^{+3} \quad \text{........................(i)}$$

Oxidation half reaction:

Atoms are already balanced.

$$H_3AsO_3 \longrightarrow H_3AsO_4$$

To balance oxygen add 1 H_2O on left side.

$$H_3AsO_3 \longrightarrow H_3AsO_4$$

To balance hydrogen add $2H^+$ on right side.

$$H_3AsO_3 + H_2O \longrightarrow H_3AsO_4 + 2H^+$$

Balance Charge:

There is no charge on the left side and 2 mono-positive charges are present on the right side corresponding to +2. Thus, right side needs $2e^-$.

$$H_3AsO_3 + H_2O \longrightarrow H_3AsO_4 + 2H^+ + 2e^- \text{.................(ii)}$$

Step 3:

Balance the loss and gain of electrons in both half reactions and add them. The reduction half-reaction has $1e^-$ while the oxidation half-reaction has $2e^-$. Thus, multiply Eq. (i) by (ii) and then add to Eq. (ii).

$$2Ce^{+4} + 2e^- \longrightarrow 2Ce^{+3}$$

$$H_3AsO_3 + H_2O \longrightarrow H_3AsO_4 + 2H^+ + 2e^-$$

$$H_3AsO_3 + H_2O + 2Ce^{+4} + 2e^- \longrightarrow H_3AsO_4 + 2H^+ + 2e^- + 2Ce^{+3}$$

Step 4: Cancel Duplication of Species on both the sides.

Duplications are 6e-. Cancel this out from both the sides.

$$H_3AsO_3 + H_2O + 2Ce^{+4} \longrightarrow H_3AsO_4 + 2H^+ + 2Ce^{+3}$$

c) $S_2O_8^{-2} + Cr^{+3} \longrightarrow SO_4^{-2} + Cr_2O_7^{-2}$

Step 1: Split the reaction into two half-reactions.

Reduction half-reaction: $S_2O_8^{-2} \longrightarrow SO_4^{-2}$

Oxidation half-reaction: $Cr^{+3} \longrightarrow Cr_2O_7^{-2}$

Step 2: Balance each half-reactions separately.

Reduction half reaction: Multiply SO_4^{-2} with 2 to balance S atoms.

$$S_2O_8^{-2} \longrightarrow 2SO_4^{-2}$$

Now O atoms are also balanced.

Balance Charge:

There is di-negative charge on the right side and there is 2 di-negative charge on the right side. Thus, left side needs $2e^-$

$$S_2O_8^{-2} + 2e^- \longrightarrow 2S_4^{-2} \text{ ....(i)}$$

Oxidation half reaction:

Balance Mass: To balance Cr atoms multiply Cr^{+3} with 2.

$$2Cr^{+3} \longrightarrow Cr_2O_7^{-2}$$

Add 7 H_2O on L.H.S. to balance O atoms.

$$2Cr^{+3} + 7H_2O \longrightarrow Cr_2O_7^{-2}$$

Add $14H^+$ on R.H.S. to balance H.

$$2Cr^{+3} + 7H_2O \longrightarrow Cr_2O_7^{-2} + 14H^+$$

Balance Charge:

There is 2 tri-positive charge on the left side corresponding to 2(+3) =+6. There is 14 mono-positive charge and di-negative charge on the right side corresponding to -2+14(+1) = +12. Thus, add $6e^-$ on R.H.S.

$$2Cr^{+3} + 7H_2O \longrightarrow Cr_2O_7^{-2} + 14H^+ + 6e^- \text{................(ii)}$$

Step 3:

Balance the loss and gain of electrons in both half reactions and add them. The reduction half-reaction has 2e⁻ while the oxidation half-reaction has 6e⁻. Thus, multiply Eq. (i) by (iii) and then add to Eq. (ii).

$$3S_2O_8^{-2} + 6e^- \longrightarrow 6SO_4^{-2}$$
$$2Cr^{+3} + 7H_2O \longrightarrow Cr_2O_7^{-2} + 14H^+ + 6e^-$$

$$2Cr^{+3} + 7H_2O + 3S_2O_8^{-2} + 6e^- \longrightarrow Cr_2O_7^{-2} + 14H^+ + 6e^- + 6SO_4^{-2}$$

Step 4:

Cancel Duplication of Species on both the sides.

Duplications are 6e⁻. Cancel this out from both the sides.

$$2Cr^{+3} + 7H_2O + 3S_2O_8^{-2} \longrightarrow Cr_2O_7^{-2} + 14H^+ + 6e^- + 6SO_4^{-2}$$

d) $MnO_4^- + C_2O_4^{-2} \longrightarrow Mn^{+2} + CO_2$

Step 1: Split te reaction into two half reactions.

Reduction half reaction: $MnO_4^- \longrightarrow Mn^{+2}$

Oxidation half reaction: $C_2O_4^{-2} \longrightarrow CO_2$

Step 2: Balance each half reaction separately.

Reduction half reaction:

Balance Mass:

Mn atoms are already balanced.

$$MnO_4^- \longrightarrow Mn^{+2}$$

There are 4 oxygen atoms on the left side and none on the right side so add 4 H_2O on left side.

$$MnO_4^- + 8H^+ \longrightarrow Mn^{+2} + 4H_2O$$

Balance Charge:

Now balance charges there are uni-negative and 8 mono-positive charges on the left side corresponding to -1+8(+1)=+7. There is di-positive charge on the right side corresponding to +2. Thus, add 5e⁻ on left side.

$$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{+2} + 4H_2O \text{ ....(i)}$$

Oxidation half reaction

Balance Mass:

Multiply CO_2 with 2 on the right side to balance C atoms

$$C_2O_4^{-2} \longrightarrow 2CO_2$$

O atoms are also balanced.

$$C_2O_4^{-2} \longrightarrow 2CO_2$$

Balance Charge:

There is di-negative charge on the left side corresponding to -2. There is no charge on the right side. Thus, add 2e⁻ on R.H.S.

$$C_2O_4^{-2} \longrightarrow 2CO_2 + 2e^- \text{ ....(ii)}$$

Step 3:

Balance the loss and gain of electrons in both half reactions and add them. The reduction half-reaction has 5e⁻ while the oxidation half-reaction has 2e⁻. Thus multiply Eq. (i) by (ii), Eq (ii) by (v) and then add the resulting equations.

$$2MnO_4^- + 16H^+ + 10e \longrightarrow 2Mn^{+2} + 8H_2O$$
$$5C_2O_4^{-2} \longrightarrow 10CO_2 + 10e^-$$

$$2MnO_4^- + 16H^+ + 5C_2O_4^{-2} \longrightarrow 2Mn^{+2} + 8H_2O + 10CO_2 + 10e^-$$

Step 4:

Cancel Duplication of Species on both the sides.

Duplications are **10e⁻**. Cancel this out from both the sides.

$$2MnO_4^- + 16H^+ + 5C_2O_4^{-2} \longrightarrow 2Mn^{+2} + 8H_2O + 10CO_2$$

Q10. Define electro-chemical cells and also give its types.

Ans: Electro-Chemical Cells:

A device, which converts electrical energy into chemical energy and vice versa, is known as electrochemical cells.

Types:

There are two types of electrochemical cells.

i. Electrolytic cells.

ii. Galvanic or voltaic cells.

Q11. What is the function of salt bridge?

Ans: The **salt bridge** allows the movement of ions from one solution to the other without mixing of the two solutions and maintains electrical neutrality in each half-cell.

Q12. Explain galvanic cell in detail. Draw a labeled diagram of a cell containing Zn-Cu electrodes.

Ans: The Galvonic Cell:

"An electrochemical cell in which spontaneous redox reaction produces an electric current is known as galvanic or voltaic cell".

The electrode at which oxidation occurs is called the **anode.** Whereas, the electrode at which the reduction occurs is called **cathode.**

Working:

When a Zn rod is dipped into a copper (II) sulphate solution, zinc atoms are oxidized to zinc ions and copper (II) ions are reduced to copper metal, which deposits on the zinc rod. Following reaction occurs:

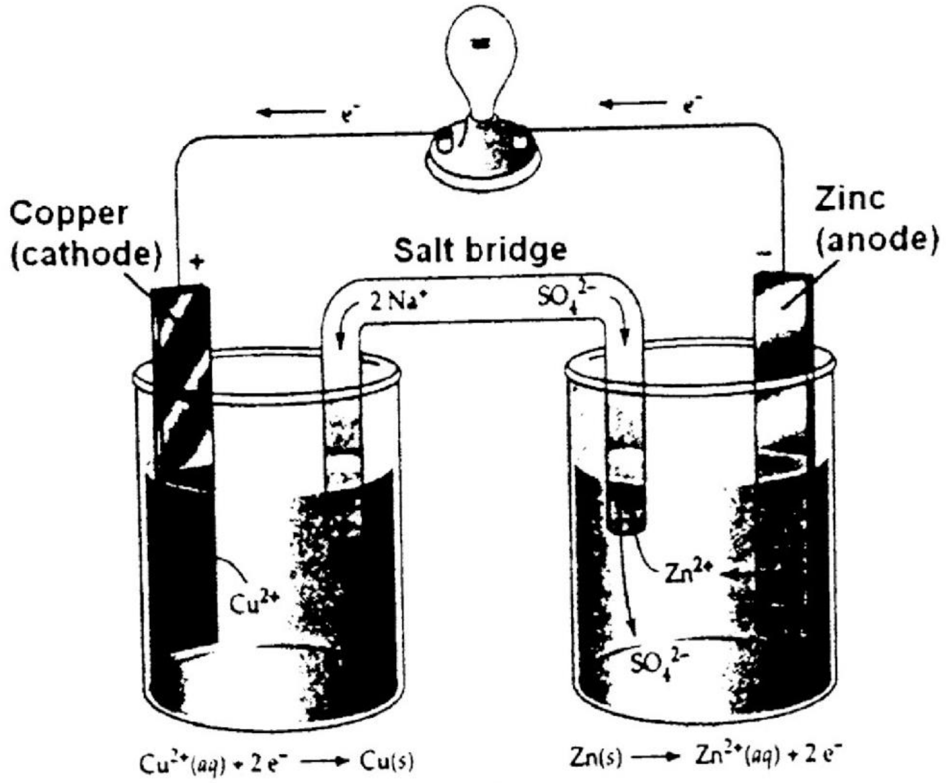
$$Zn(s) + Cu^{+2}(aq) \longrightarrow Zn^{+2}(aq) + Cu(s)$$

In this reaction, electrons flow directly from the zinc rod to Cu^{+2} ions in solution. However, if the electrons transfer from Zinc rod to the copper ions in solution could be directed through an external circuit, the spontaneous redox reaction could be used to generate electric current.

Construction:

A zinc rod is dipped in zinc sulphate solution in one container is connected by a copper wire with the copper metal and the copper rod dipped in copper (II) sulphate solution in a separate container, no current flows through the external circuit.

However, when the two solutions are connected with a tube (**salt bridge**) filled with a solution of an electrolyte such as KCl, KNO_3 or Na_2SO_4, current flows through external circuit.



A simple Galvanic Cell

Oxidation half reaction:

In one half-cell oxidation takes place and is called as oxidation half-cell or **anode**. Reaction taking place in oxidation half-cell is called oxidation half reaction and

Reduction half reaction:

In the other half-cell reduction takes place is called as reduction half-cell or **cathode**. The reaction taking place in reduction half-cell is called reduction half reaction.

Flow of electron:

Zn has greater tendency to lose electrons than Cu. Therefore, Zn electrode acquires negative charge relative to Cu electrode. The electrons flow from Zn electrode through the external circuit to Cu electrode.

Reactions occur at electrodes:

The following half-cell reactions occur at the two electrodes.

At anode: $Zn\ (s) \longrightarrow Zn^{+2}(aq) + 2e^-$ (Oxidation half reaction)

At cathode: $Cu^{+2}(aq) + 2e^- \longrightarrow Cu\ (s)$ (Reduction half reaction)

Over all cell reaction: $Zn\ (s) + Cu^{+2}\ (aq) \longrightarrow Zn^{+2}(aq) + Cu(s)$

Q13. What is Cell Potential?

Ans: Cell Potential:

The force with which electrons are pushed to flow through the wire from anode to cathode is called the **electromotive force or emf**.

The emf produced by galvanic cell is called cell potential (E^0 cell). It depends upon the difference in the electrode potentials of the two half cells joined in series. It is measured is volts (V).

Volt:

The volt is the measure of energy that is capable of being extracted from the flowing electric charge. When the passage of one coulomb is able to accomplish one Joule of work, then emf is one volt.

Explanation:

i. The electrode with the more negative reduction potential acts as anode and the electrode with the more positive reduction potential acts as cathode.

ii. The net reaction that occurs in the galvanic cell under standard conditions can be constructed from standard reduction half-reactions. Since a redox reaction must include both an oxidation reaction and a reduction reaction, one of the half-cells must run an oxidation that supplies electrons for the reduction.

iii. Electrons always flow spontaneously from more negative electrical potential to more positive electrical potential. This means that, under standard conditions, electrons are produced by the reaction with the more negative standard potential and consumed by the reaction with the more positive standard potential. Thus under standard conditions (1 mol dm^{-3} concentration at 25 °C and 1 atm pressure), the reaction with more negative E^0 value occurs as oxidation (anode reaction). The reaction with more positive E^0 value occurs as reduction (cathode reaction).

iv. The voltage of any cell under standard conditions can be calculated using standard reduction potentials.

v. Any combination of two half-cells will produce a complete cell. The overall cell reaction is obtained by suitably combining the equations for the two half reactions.

vi. Standard cell potential E°cell or emf of cell is the algebraic difference between the respective standard reduction potentials of the two half-cells.

$$E^0_{cell} = E^0_{cathode} - E^0_{anode}$$

The cell potential has positive value for any spontaneous redox reaction.

Measurement of electrode potential:

The potential of a galvanic cell can be measured with a voltmeter. But a single half-cell potential or the **electrode potential** cannot be measured directly. This is because one half-cell reaction cannot occur without a simultaneous reaction in another half-cell.

However, relative half-cell potential (electrode potential) can be determined by arbitrarily designating one half-cell reaction and its potential as the standard to which other half-cell potentials are compared. By international agreement, this reference half-cell is the **standard hydrogen electrode**, with the standard potential of 0.00V.

Q14. Define standard electrode potential.

The standard electrode potential is defined as the tendency of a half-cell reaction to undergo reduction relative to the standard hydrogen electrode.

Explanation:

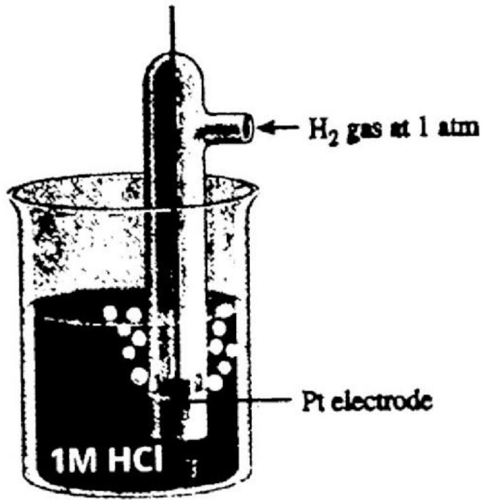
It is potential difference developed when an electrode of an element is placed in a solution containing ions of that element when all the components are in their standard state i.e. 1 atm for gases, 1M for solutions, pure solid for electrode and at 25°C.

Q15. Describe standard hydrogen electrode (SHE) also write the potential for oxidation and reduction reactions.

Ans: Standard Hydrogen Electorde:

Construction:

A standard hydrogen electrode (SHE) consists of a platinum foil coated with finely divided platinum, surrounded by hydrogen gas at 1atm pressure in contact with IM HCl solution at 298K. Its electrode potential is arbitrarily chosen as zero at all temperatures.



Standard Hydrogen Electrode

By convention, the half-cell potential for reduction of $H^+(aq)$ to H_2 gas or the potential for the oxidation of H_2 to $H^+(aq)$ in the standard Hydrogen half-cell is defined exactly 0.00V.

Redox reaction:

Reduction: (act as cathode):

$$2H^+(aq) + 2e^- \longrightarrow H_2(g) \qquad E^0_{H^+/H_2} = 0.00V$$

Oxidation: (act as anode):

$$H_2(g) \longrightarrow 2H^+(aq) + 2e^- \qquad E^0_{H_2/H^+} = 0.00V$$

- The symbol E^0 designates a standard potential i.e. the potential measured under standard conditions (1M concentration, 1 atm pressure and 25°C).

Table: Reduction potentials of some elements, ions and compounds:

Reduction Half-reaction	E0 (Volts)
$Li^+ + \bar{e} \rightleftharpoons Li$	–3.05
$K^+ + \bar{e} \rightleftharpoons K$	–2392
$Ba^{2+} + 2\bar{e} \rightleftharpoons Ba$	–2.90
$Ca^{2+} + \bar{e} \rightleftharpoons Na$	–2.76
$Na^{2+} + 2\bar{e} \rightleftharpoons Na$	–2.71
$Mg^{2+} + 2\bar{e} \rightleftharpoons Mn$	–2.38
$Al^{3+} + 3\bar{e} \rightleftharpoons Al$	–1.67
$Mn^{2+} + 2\bar{e} \rightleftharpoons Mn$	–1.03
$2H_2O + 2\bar{e} \rightleftharpoons H_2 \because 2OH^-$	–0.83
$Zn^{2+} + 2\bar{e} \rightleftharpoons Zn$	–0.76
$Cr^{3+} + 3\bar{e} \rightleftharpoons Cr$	–0.74

$Fe^{2+} + 2\bar{e} \rightleftharpoons Fe$	–0.44
$PbSO_4 + 2\bar{e} \rightleftharpoons Pb + SO_4^{2-}$	–0.36
$Ni^{2+} + 2\bar{e} \rightleftharpoons Ni$	-0.25
$Sn^{2+} + 2\bar{e} \rightleftharpoons Sn$	-0.14
$Pb^{2+} + 2\bar{e} \rightleftharpoons Pb$	-0.13
$Fe^{3+} + 3\bar{e} \rightleftharpoons Fe$	-0.04
$2H^{+} + 2\bar{e} \rightleftharpoons H_2$	0.00
$AgCl + \bar{e} \rightleftharpoons Ag + Cl^-$	+0.22
$Hg_2Cl_2 + 2\bar{e} \rightleftharpoons 2Hg + 2Cl^-$	+0.27
$Cu^{2+} + 2\bar{e} \rightleftharpoons Cu$	+0.34
$Cu^{+} + 1\bar{e} \rightleftharpoons Cu$	+0.52
$I_{2(aq)} + 2\bar{e} \rightleftharpoons 2I^-$	+0.54
$Fe^{3+} + \bar{e} \rightleftharpoons Fe^{2+}$	+0.77
$Ag^{+} + \bar{e} \rightleftharpoons Ag$	+0.80
$Br_{(aq)} + 2\bar{e} \rightleftharpoons 2Br^-$	+1.09
$O_2 + 4H+ + 4\bar{e} \rightleftharpoons 2H_2O$	+1.23
$MnO_2 + 4H+ + 2\bar{e} \rightleftharpoons Mn^{2+} + 2H_2O$	+1.28
$Cr_2O_7^{2-} + 14H + + 6\bar{e} \rightleftharpoons 2Cr^{3+} + 7H_2O$	+1.33
$Cl_2\ (g) + 2\bar{e} \rightleftharpoons 2Cl^-$	+1.36
$2ClO_3 + 12H+ + 5\bar{e} \rightleftharpoons Cl_2 + 6H_2O$	+1.47
$8H^{+} + MnO_4^- + 5\bar{e} \rightleftharpoons Mn^{2+} + 4H_2O$	+1.49
$PbO_2 + SO_4^{-2} + 4H^{+} + 4\bar{e} \rightleftharpoons PbSO_4 + 2H_2O$	+1.69
$H_2O_2 + 2H+ + 2\bar{e} \rightleftharpoons 2H_2O$	+1.7
$S_2O_3^{-2} + 2\bar{e} \rightleftharpoons 2SO_4^{-2}$	+2.00
$F_2 + 2\bar{e} \rightleftharpoons 2F^-$	+2.87

Q16. How can we determine cell potential? Explain your answer with the help of example.

Ans: Determination of Cell Potential:

A cell reaction consists of two half reactions.

Reduction takes place in the half-cell having greater value of reduction potential.

Oxidation takes place in the half-cell having the smaller value of reduction potential.

Equation of the half-cell reaction having smaller value of reduction potential is reversed and added to the equation of half-cell having greater value of reduction potential. Sum of these two equations represent cell reaction.

Example:

Calculate E^0 cell for Zn-Cu cell and write cell reactions. Show direction of electron flow.

Solution:

Half Cell reaction			**Reduction potential**
i)	$Zn^{++}\ (aq) + 2\bar{e} \longrightarrow$	$Zn\ (s)$	$-0.76V$
ii)	$Cu^{++}\ (aq) + 2\bar{e} \longrightarrow$	$Cu\ (s)$	$+0.34V$

Data indicates that the reduction potential of second half-cell is greater than the first. Hence reduction reaction will occur in second half-cell and oxidation in the first half-cell. Reverse the first equation and add it to the second equation to get cell reaction.

$$Zn(s) \longrightarrow Zn^{+2}(aq) + 2\bar{e} \quad \text{(Oxidation half-reaction)}$$
$$Cu^{+2}(aq) + 2\bar{e} \longrightarrow Cu(s) \quad \text{(Reduction half-reaction)}$$

$$Zn(s) + Cu^{+2}(aq) \longrightarrow Cu(s) + Zn^{+2}(aq) \quad \text{(Cell reaction)}$$

$$E^0_{cell} = E^0_{cathode} - E^0_{anode}$$
$$E^0_{cell} = E^0_{Cu} - E^0_{Zn}$$
$$E^0_{cell} = +0.34 - (-0.76)$$
$$E^0_{cell} = +1.10V$$

Electrons will flow from anode to cathode i.e., from Zn electrode to Cu electrode.

Example 12.11:

The standard reduction potentials for the following half reactions are:

$$Ni^{+2}(aq) + 2e^- \longrightarrow Ni(s) \quad E^0 = -0.25V$$
$$Mg^{+2}(aq) + 2e^- \longrightarrow Mg(s) \quad E^0 = -2.38V$$

Calculate E^0_{cell} for Ni-Mg cell, write cell reactions, show direction of electron flow and identify the anode of the cell.

Solution:

Data indicates that the reduction potential of first reaction is greater than that of the second reaction. Hence reduction will occur in the first reaction and oxidation in the second reaction. Reverse the second reaction and add it to the first reaction to get the cell reaction.

$$Mg(s) \longrightarrow Mg^{+2}(aq) + 2e^- \quad \text{(Anode reaction)}$$
$$Ni^{+2}(aq) + 2e^- \longrightarrow Ni(s) \quad \text{(Cathode reaction)}$$

$$Mg(s) + Ni^{+2}(aq) \longrightarrow Mg^{+2}(aq) + Ni(s) \quad \text{(Cell reaction)}$$

Thus Mg will act as anode and Ni as cathode. Electrons will flow from Mg to Ni.

$$E^0_{cell} = E^0_{cathode} - E^0_{anode}$$
$$E^0_{cell} = E^0_{Ni} - E^0_{Mg}$$
$$= -0.25 - (-2.38) = 2.13\ V$$

SELF-CHECK EXERCISE 12.7

The standard reduction potentials for the following half-reactions are:

$$Cu^{+2}(aq) + 2e^- \longrightarrow Cu(s) \quad E^0 = +0.34V$$
$$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq) \quad E^0 = +1.36V$$

Calculate E^0_{cell} for Cu-Cl_2 cell, write cell reactions, identify cathode and show the direction of electron flow.

Solution:

Given that the reduction potential of second reaction is greater than that of first reaction. Therefore reduction will occur in the second reaction and oxidation will occur in the first reaction. To get the cell reaction reverse the first reaction and add it to the second reaction to get the cell reaction.

Anode Reaction: $Cu\,(s) \rightarrow Cu^{+2}(aq) + 2e^-$ (Oxidation half-reaction)
Cathode Reaction: $Cl_2\,(g) + 2e^- \rightarrow 2Cl^-(aq)$ (Reduction half-reaction)

Cell reaction: $Cu\,(s) + Ni^{+2}(aq) \rightarrow Cu^{+2}(aq) + 2Cl^-(aq)$ (Redox Reaction)

Thus Cu will act as anode and Cl_2 as cathode. Electrons will flow from Cu to Cl_2. The cell potential is

$$E^0_{\text{cell}} = E^0_{\text{cathode}} - E^0_{\text{anode}}$$
$$E^0_{\text{cell}} = E^0_{Cl_2} - E^0_{Cu}$$
$$= +1.36 - (+0.34) = +1.02V$$

Q17. How can we determine the feasibility of a chemical reaction?

Ans: Feasibility of a Chemical Reaction:

The feasibility of a chemical reaction can be determined by the sign of the sum of E^0 values of the two half-cell reactions.

Condition:

i. The positive value indicates that the reaction occurs spontaneously or will be feasible.

ii. The negative value indicates that the reaction is not feasible.

Example 12.12:

Is the following reaction feasible?

$$Sn + Fe^{+2} \longrightarrow Sn^{+2} + Fe$$

The standard reduction potential values are $E^0_{Sn} = -0.14V$, $E^0_{Fe} = -0.44V$

Solution:

Oxidation

$$Sn^0 + Fe^{+2} \longrightarrow Sn^{+2} + Fe^0$$

Reduction

It is clear from the above equation that oxidation of Sn and reduction of Fe is taking place. Sn is acting as anode and Fe as cathode. The above reaction consists of the following two half-cell reactions.

$Sn \longrightarrow Sn^{+2} + 2\bar{e}$ (oxidation half reaction)

$Fe^{+2} + 2\bar{e} \longrightarrow Fe$ (reduction half reaction)

$$E^0_{\text{cell}} = E^0_{\text{cathode}} - E^0_{\text{anode}}$$
$$E^0_{\text{cell}} = -0.44V - (-0.14V) = -0.30$$

As E^0_{cell} is negative, therefore the given reaction is not feasible. However reverse reaction would be spontaneous.

SELF-CHECK EXERCISE 12.8

Using emf data, explain the following:

1. **Can Fe displace Cu from a solution of Copper (II) Sulphate.**
2. **Can Iodine displace Bromine from aqueous solution of Potassium bromide?**

Solution:

1. E^0_{Cu} = 0.34 V

E^0_{Fe} = -0.448 V

$$E^0_{cell} = E^0_{cathode} - E^0_{anode}$$

$$E^0_{cell} = E^0_{Cu} - E^0_{Fe}$$

$$E^0_{cell} = +0.34V - (-0.44V) = +0.78V$$

As E^0_{cell} is positive, therefore the reaction is feasible. Fe ranked above Cu in electrochemical series therefore it can displace Cu^{2+} from a solution of copper (II) sulphate.

2. E^0_{I} = 0.54 V

E^0_{Br} = 1.09 V

$$E^0_{cell} = E^0_{cathode} - E^0_{anode}$$

$$E^0_{cell} = E^0_{I} - E^0_{Br}$$

$$E^0_{cell} = +0.54V - (+1.09V) = -0.55V$$

As E^0_{cell} is negative, therefore the reaction is not feasible. Hence, iodine cannot displace Bromine from aqueous solution of potassium iodide.

Important Information

In modern dentistry, a material most commonly used to fill decaying teeth is known as dental amalgam. Dental amalgam actually consists of Ag, Sn and Hg, the standard electrode potential for this amalgam is +0.67V. Any person who bites a piece of Al foil (such as use for wrapping candies, biscuits etc) in such a way that the foil presses against a dental filling, will experience a momentarily sharp pain. This is because, an electro chemical cell has been created in the mouth, with Al as anode E°= −1.66V, the filling as cathode and saliva as electrolyte. Contact between Al foil and filling short circuits the cell. This causes the weak current to flow between the electrodes. This current stimulates the sensitive nerve of the tooth, causing a sharp pain.

Q18. What are electrochemical series?

Ans: Electrochemical Series:

Under the recommendation of international union of pure and applied chemistry (IUPAC) the half-cell reactions are given in the reduction reactions therefore E° values are known as reduction potentials.

However, the value of oxidation potential for an electrode can be obtained by reversing the sign of reduction potential for that electrode.

Note: The given reduction potential values relate to standard conditions only. i.e., 1M solution of ions, 25°C (298K) and 1 atm pressure. Changes in conditions will alter these values. Such a list of arrangement of elements in the order of their standard electrode potential with reference to standard hydrogen electrode is called **electrochemical series.**

Q19. Explain the activity series of metals with the help of examples also draw the chart which shows the activity of metals.

Ans: Activity Series of Metals:

Metals are ranked according to their ability to replace other metals and hydrogen from their compounds.

In this ranking metals and hydrogen are arranged in order of decreasing ease of oxidation to their respective ions in aqueous solution.

This arrangement is called **activity series**.

Explanation:

A displacement or replacement reaction occurs when an element displaces another element which is a part of a compound.

General equation: $A + XY \longrightarrow X + AY$

In the general equation atom A replaces atom X and the compound XY illustrate this type of reaction.

Example:

1. If Zn metal is placed in a blue solution of copper (II) sulphate, the blue color slowly fades away and grey metal is replaced by red orange Cu metal. In this reaction Cu ions in the solution are reduced to Cu metal and Zn atoms are oxidized to Zn ions.

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

When copper metal is placed in zinc sulphate solution, no replacement reaction occurs.

Reason:

As the standard reduction potential reaction of copper is greater than that of zinc therefore no replacement reaction occurs.

Conclusion:

- **i.** This means that it is easy to oxidize Zn to its ions and reduce Cu^{+2} ions to its atoms. Thus Zn can replace Cu^{+2} ions from its solution.
- **ii.** That is why Mg and Al can also displace Cu^{+2} ions, but Ag cannot replace Cu^{+2} ions.
- **iii.** It is observed that metal like Na, K can displace H_2 from water but metals like Cu, Ag cannot displace H_2 from water.

Chart of activity series of common metals:

Li
K
Ba
Sr
Ca
Na

Very Active metals;
react with cold water with the liberation of hydrogen gas;
(K and Na react violently with water),
they also react violently with acids

Mg
Al
Mn
Zn
Cr
Fe
Cd

Metals of intermediate activity:
React with steam or with acids such as HCl with liberation of H_2

Co
Ni
Sn
Pb
} Moderately active metals
React slowly with HCl
Do not react with water

H_2

Cu } Moderately noble metals;
Ag } do not react with water, HCl but react with oxidizing acids such as HNO_3
Hg } and $HClO_4$

Pt } Very noble metals;
Au } react only with aqua regia

Q20. Write important features of activity series of common metals.

Ans: Important features activity series are as follows:

1. Metals high on the list transfer electrons to metal cations lower on the list. The greater the separation between the species the more vigorous will be the reaction.

Example:

When powdered barium is heated with lead (II) oxide, a replacement reaction occurs. This is because Ba is more active than Pb. Barium is oxidized to form barium oxide and lead is reduced to elemental lead.

$$Ba\ (s) + PbO\ (s) \longrightarrow BaO\ (s) + Pb\ (s)$$

On the other hand, when iron pellets are added to a solution of $MgCl_2$ no reaction will occur. This is because Fe is below Mg in the activity series.

2. Very active metals react with cold water to liberate Hydrogen.

Example:

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

3. The less active metals react with steam and with non-oxidizing acid such as HCl.

Example:

$$Mg\ (s) + 2H_2O\ (g) \longrightarrow Mg(OH)_2\ (aq) + H_2\ (g)$$
$$Mg\ (s) + HCl\ (aq) \longrightarrow MgCl_2\ (aq) + H_2\ (g)$$

Al reacts with hot water to a small degree. This is because Al forms a protective coating of Aluminum hydroxide and Aluminum oxide on its exposed surface. This protects metal from further reaction.

4. The moderately active metals such as CO, Ni, Sn, Pb do not react with steam. These metals react spontaneously with 1M HCl.

5. The metals below hydrogen in the activity series do not react with 1 M HCl. These metals are unable to reduce H^+ ions in 1M HCl solution. However, their 1M aqueous ions can be reduced by hydrogen gas at 1 atm. These metals are called as noble metals.

The moderately noble metals Cu, Ag and Hg react only with oxidizing acids such as nitric acid and perchloric acids. The very noble metals do no react with these acids but reaction with aqua regia.

Example:

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(g)$$

Example 12.13:

Predict whether a replacement reaction will occur in the following instances. Explain your conclusion.

a) Magnesium ribbon is in a solution of silver nitrate.

b) A small piece of calcium is added to a beaker of water.

c) A copper wire is dipped in 1M HCl.

Solution:

a) Magnesium is above Silver in the activity series. Thus Mg will displace Ag^+.

$$Mg\ (s) + 2Ag^+\ (aq) \longrightarrow Mg^{+2}\ (aq) + 2\ Ag\ (s)$$

b) Calcium being very active metal and above hydrogen in the activity list will react with cold water and will liberate hydrogen gas.

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

c) Copper metal is below hydrogen in the activity list. therefore no reaction will take place.

SELF-CHECK EXERCISE 12.9

Predict whether a reaction occurs in the following cases and write a net ionic equation for the reactions that occurs:

a) An iron nail is placed in 1M HCl.

b) Lead (II) oxide is heated with powdered zinc.

c) Nickel wire is place into a solution of silver nitrate.

Solution:

a) Iron is ranked higher in activity series than hydrogen. Thus it can displace hydrogen from $HCl$and the following reaction will occur.

$$2Fe\ +\ 6HCl\ \rightarrow\ 2FeCl_3\ +\ 3H_2$$

b) Zinc is ranked higher in activity series than lead. Thus it can displace lead from Lead (II) oxide and the following reaction will occur.

$$Zn\ +\ PbO\ \longrightarrow\ ZnO\ +\ Pb$$

c) Nickel is ranked higher in activity series than silver. Thus it can displace silver from silver nitrate and the following reaction will occur.

$$Ni\ +\ 2AgNO_3\ \rightarrow\ Ni(NO_3)_2\ +\ 2Ag$$

Q21. How would you explain electrolytic cells in terms of electrolysis?

Ans: Electrolytic Cells:

In these cells electrical energy is used to drive many chemical processes.

Example:

For example heavy industrial processes such as the preparation of sodium hydroxide, metals, purification of nickel and copper, plating of noble metal on jewellery and instruments.

Electrolysis:

The chemical process used in electrolytic Cells is called electrolysis.

OR

Electrolysis is the production of a chemical reaction by means of an electric current. The apparatus use for electrolysis consists of an electrolytic cell.

Construction:

Electrolytic cell contains the electrolyte either in molten state or in solution into which two electrodes are placed.

The electrodes are connected with a battery.

The current is carried from the battery through the wires by mean of electrons (metallic conduction).

Working:

i. Within the cell the current is carried by the anions and cations of the electrolyte (Electrolytic conduction).

ii. The electrodes serve as a point where conduction changes from metallic to electrolytic or vice versa.

iii. At each electrode a chemical reaction takes place in which electrons are gained by the ions in solution at one electrode.

iv. Simultaneously electrons are released by some substance at the other electrode.

v. These electrons are returned to the battery through the connecting wire.

vi. Thus oxidation-reduction reactions occur at the electrodes.

vii. The electrode at which oxidation occurs is called as anode.

viii. The electrode at which reduction occurs is called as cathode.

Note: The changes, which occur at the electrodes, depend on the relative oxidation–reduction tendencies of the substances involved.

Q22. Define the following terms.

a. Coulomb (C) b. Faraday (F) c. Ampere

Ans: a. Coulomb (C):

The SI unit of charge is the **coulomb (C)**. It is the charge on 6.25×10^{18} electrons.

b. Faraday (F):

Although the coulomb is the usual unit for measuring charge, the chemist finds that a more convenient unit is the **Faraday (F)**. It corresponds to the charge carried by the mole of electrons and it is equal to 96487 C.

c. Ampere:

The SI Unit of current is the ampere, which is the amount of current flowing when one coulomb passes a given point in one second. Frequently, an ampere is referred as "a coulomb per second".

Q23. How can you relate chemical change and electric current in terms of Faraday's laws?

Ans: Relation Between Chemical Change And Electric Current:

In 1833, Faraday described the results of his electrochemical investigations by stating the two principles of electrochemistry, which are now known as **Faraday's laws**.

Faraday's first law:

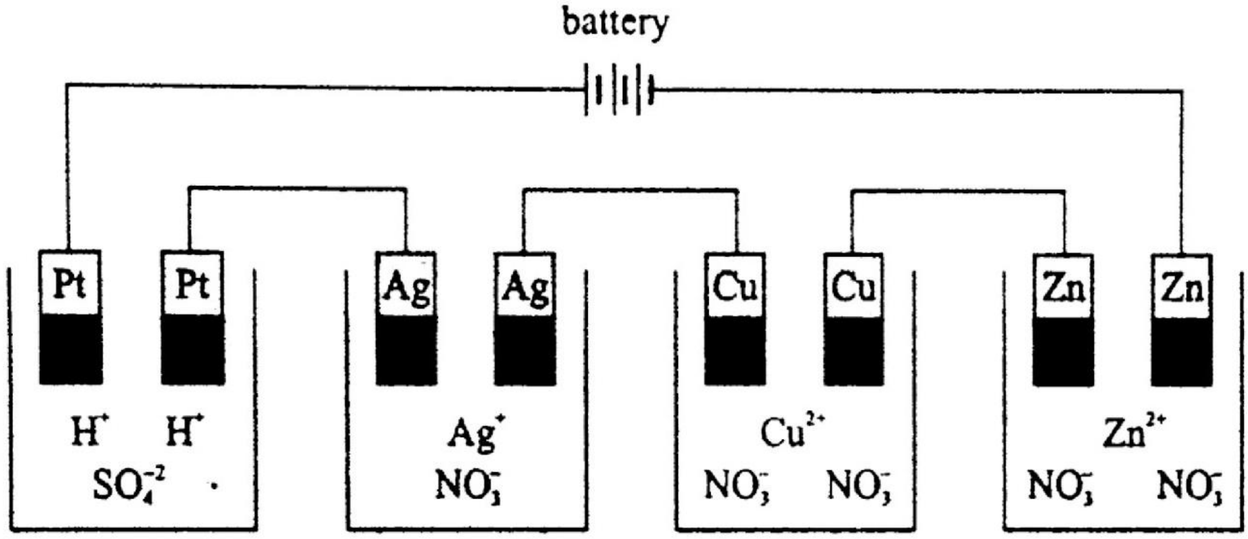
The first law of Faraday states that "**The amount of chemical reaction taking place at an electrode is directly proportional to the quantity of charge that flows through the electrode during the process.**"

Faraday's second law:

Second law of Faraday states that. "**If same quantity of charge is passed through different electrolytic cells, the chemical change in each case is proportional to the gram equivalent mass of each parent species**".

Explanation:

If several cells containing aqueous solutions are connected in series, as shown in figure and if 96487 C of charge is passed through them, the electrode reactions proceed simultaneously and for 7.999g of O_2 produced from H_2SO_4, 1.008g of H_2, 107.9g of Ag, 31.77g of Cu and 32.5g of Zn are produced at the respective cathodes. These weights are the equivalent weights of the respective elements.



Electrolysis of Aqueous Solutions

It is therefore concluded that 96487 C is the charge on one mole of electrons. This quantity of charge is referred as one Faraday. Thus the quantity of the change that occurs in electrolysis can be determined from the number of Faraday's of charge, which passes. For most calculations, the value of the Faraday will be taken as 96500 C.

Conclusion:

The amount of a substance produced during electrolysis by passing one Faraday of electricity is called its equivalent weight.

Example 12.14:

In the electrolysis of molten $ZnCl_2$, how much Zn can be deposited at the cathode by passage of 0.01 ampere for one hour?

Solution:

0.01 Amp. for one hour $= 0.01 \times 1 \times 60 \times 60 = 36$ C

96500C/Faraday $= 3.7 \times 10^{-4}$ Faraday

In molten Zinc chloride, the cathode reaction is

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

Which means that for every 2 Faraday of electricity used up, one mole of Zn is deposited. Thus,

3.7×10^{-4} Faraday × 1 Mole of Zn / 2 Faradays $= 1.85 \times 10^{-4}$ mole of Zn
As one mole of Zn is 63.37g.
Therefore 1.85×10^{-4} mole of Zn × 63.37g = 0.012g of Zn (Ans)

Example 12.15:

A constant current was passed through a solution of $AuCl_4^-$ ions between gold electrodes. After a period of 10.0 minutes, the cathode increased in weight by 1.314 grams.

i) How much charge was passed?
ii) What was the amount of current?
iii) What volume of Cl_2 was collected at anode at 1 atm and 25°C.

Solution:

The reaction at cathode is the reduction of Au (III) to Au metal

$$AuCl_4^- + 3e^- \longrightarrow Au + 4Cl^-$$

It means that for every 3 Faraday of electricity used up, 1 mole of Au is produced.

Moles of Au $= \dfrac{1.314g\ Au}{197g/mole\ of\ Au} = 6.67\times10^{-3}$

i) Charge $= 6.67\times10^{-3}$ mole Au $\times \dfrac{3\ Faraday}{mole\ of\ Au}$

$= 2\times 10^{-2}$ Faraday

ii) Current $= \dfrac{Charge}{Time(s)}$

Time = 10 mins = 10 × 60 = 600s

Current $= \dfrac{(2\times10^{-2}F)(96500C/F)}{600s} = 3.22$ A.

iii) The reaction at anode is oxidation of Cl^- ions.

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

For every 2 Faraday of electricity 1 mole of Cl_2 was produced.
For 1 Faraday of electricity = 1/2 moles of Cl_2 was produced.
For 2×10^{-2} Faraday of electricity = $1/2 \times 2\times10^{-2}$ moles of Cl_2 was produced.
= 1×10^{-2} moles of Cl_2 was produced.
Volume of Cl_2 produced can be calculate by the following formula:

$$V = \frac{nRT}{P} = \frac{1 \times 10^{-2} \times 0.08205 \times 298}{1} = 0.245\ dm^{-3}$$

SELF-CHECK EXERCISE 12.10

1. **Bauxite ore is used for the commercial preparation of Al. For this purpose bauxite ore is first purified to produced pure alumina, Al_2O_3. Alumina is then electrolyzed. Following reaction occurs:**

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

Calculate mass of Al, that collects at the cathode and volume of oxygen that collects at node when Al_2O_3 is electrolyzed for 10 hours with a 15 ampere current at 1 atm and 25°C.

Solution:

Temperature = 25 °C + 273 = 298

Pressure = 1 atm

General gas constant = R = 0.0821 atm dm^3 Mole^{-1} K^{-1}

Time = 10 hours = 10 × 60 × 60 = 36000s

Current = 15 A

$$Current = \frac{Charge}{Time\ (s)}$$

$$Charge = Current \times Time\ (s)$$
$$= 15A \times 36000\ (s) = 540000\ C$$

Mass of Al:

In the electrolysis of Al_2O_3, the cathode reaction is

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

It indicate that for every 3 moles of electrons, 1 mole of Al is produced.

1F = 1 mole of electron.

3F = 3 moles of electrons = 1 mole of Al

So,

3F charge produces Al = 1 mole

5.596 F charge produces Al $= \frac{1}{3} \times 5.596 = 1.865\ moles$

$Atomic\ mass\ of\ Al = 27g\ mole^-$

$1\ mole\ of\ Al = 27\ g$

1.865 moles of Al = 27 × 1.865 = $50.355\ g$

Volume of O_2:

Oxidation of O^{2-} ions occurs at anode.

$$2O^{2-} \longrightarrow O_2 + 4e^-$$

Thus for every 4 moles of electrons. 1 mole of O_2 is produce.

1F = 1 mole of electron

4F = 4 moles of electrons = 1 mole of O_2

4 F electricity produce O_2 = 1 mole

5.596 F of electricity produce $O_2 = \frac{1}{4} \times 5.596 = 1.399\ moles\ of\ O_2$

Number of moles of O_2 = n = $1.399\ moles$

Volume = V = ?

$$PV = nRT$$
$$V = \frac{nRT}{P}$$
$$V = \frac{1.399 \times 0.0821 \times 298}{1} = 34.228\ dm^3$$

2. Which of the following compounds will give more mass of metal, when 15 ampere current is passed through molten mass of these salts for 1 hr. a) NaCl (b) $CaCl_2$

Solution:

Time = 1 hours = 1 ×60 × 60 = $3600\ s$

Current = 15A

$$Current = \frac{Charge}{Time\,(s)}$$

$$Charge = Current = Time\,(s)$$

$$= 15A \times 3600\,s = 54000\,C$$

$$As\ 96500\,C = 1F$$

$$54000\,C = \frac{1F}{96500\,C} \times 54000\,C = 0.5596\,F$$

(a) NaCl:

In the electrolysis of molten $NaCl$, the cathode reaction is

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

For every 1 mole of electrons, 1 mole of Na is produced.

Since, 1 F = 1 mole of electron = 1 mole of Na

1 F charge produces Na = 1 mole

$$0.5596\ F\ charge\ produces\ Na = \frac{1}{1} \times 0.5596 = 0.5596\ moles$$

Atomic mass of Na = 23 g $mole^{-1}$

$1\ mole\ of\ Na = 23\ g$

$0.5596\ moles\ of\ Na = 23 \times 0.5596 = 12.871\ g$

(b) $CaCl_2$

In the electrolysis of molten $CaCl_2$, the cathode reaction is

$$Ca^{+1} + 2e^{-} \longrightarrow Ca$$

For every 2 mole of electrons, 1 mole of Ca is produced.

1 F = 1 mole of electron

2F = 2 moles of electrons = 1 mole of Ca

2F charge produces Ca = 1 mole

$$0.5596\ F\ charge\ produces\ Na = \frac{1}{2} \times 0.5596 = 0.2798\ moles$$

Atomic mass of Ca = 40 g $mole^{-1}$

$1\ mole\ of\ Ca = 40\ g$

$0.2798\ moles\ of\ Ca = 40 \times 0.2798 = 11.192\ g$

Thus $NaCl$ will produce more mass of Na

Q24. What are batteries? Also give their applications.

Ans: Batteries:

A battery is a galvanic cell or a group of galvanic cell connected in series.

Batteries are source of direct current and have become essential source of portable power in our society. A battery can be as tiny as a heart pacemaker implant or as large as the charge storage tanks of an electric automobile.

Applications:

Batteries provide electric power for starting internal combustion engines in automobiles, for running systems on space vehicles and for such devices as flash lights, toys, heart pacers, electronic calculators, portable radios, TVs, Tape recorders etc.

Q25. Define primary cells and secondary cells.

Ans: Primary cells:

Batteries which cannot be recharged are called primary cells e.g. dry cell.

Secondary cells:

The batteries which can be recharged are known as secondary cells or storage batteries, e.g. Lead storage battery (automobile battery).

Q26. Explain dry cell, give its types and also write the redox reaction taking place at anode and cathode.

Ans: Dry Cell:

The dry cell batteries are used to power many flashlights, toys and small appliances.

Construction:

The anode [illegible] cathode is an inert graphite rod at the center of the container in contract with a mixture of MnO_2 and carbon (charcoal).

Electrolyte: The electrolyte is a mixture of moist NH_4Cl and $ZnCl_2$

Reactions: Following reactions take place in it.

At Anode: $Zn \longrightarrow Zn^{+2} + 2\bar{e}$

At Cathode: $2NH_4^+ + 2MnO_2 + 2\bar{e} \longrightarrow Mn_2O_3 + 2NH_3 + H_2O$

This cell produces a **potential of 1.5V**.

Types:

i. Alkaline dry cell:

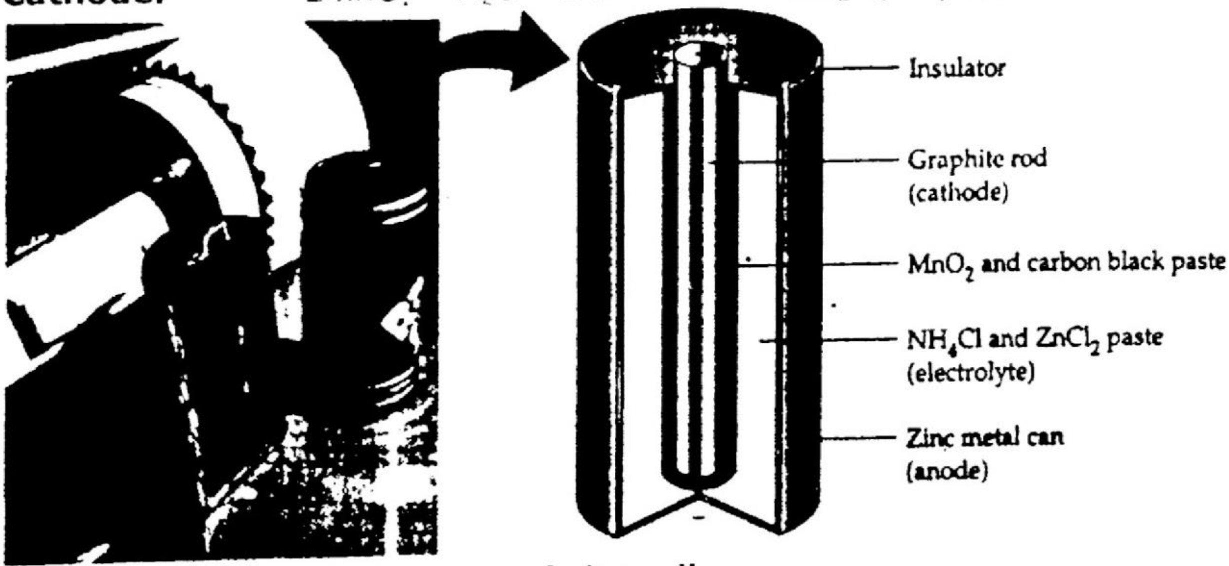
Construction:

In the alkaline dry cell battery, moist paste of KOH is used as electrolyte instead of NH_4Cl and $ZnCl_2$

Reactions: Following reactions take place in it.

At anode: $Zn + 2OH^- \longrightarrow Zn(OH)_2 + 2\bar{e}$

At Cathode: $2MnO_2 + H_2O + 2\bar{e} \longrightarrow Mn_2O_3 + 2OH^-$



A dry cell

Alkaline dry cell lasts longer:

The alkaline dry cell lasts longer because the zinc anode corrodes less rapidly in basic conditions.

ii. Nickel-cadium dry cell:

An important type of dry cell is the nickel-cadium battery

Construction:

It has a Cd anode and NiO_2 as cathode. KOH is an electrolyte.

Reactions: Following reactions occur in it.

At anode: $Cd\,(s) + OH^-\,(aq) \longrightarrow Cd\,(OH)_2\,(s) + 2e$

At cathode: $NiO_2\,(s) + 2H_2O\,(l) + 2e^- \longrightarrow Ni(OH)_2\,(aq) + 2OH^-\,(aq)$

Recharging:

In this cell, the products adhere to the electrodes. Thus, battery can be recharged.

Q27. What do you know about lead storage battery? Write the redox reactions taking place at anode and cathode during discharging and recharging of lead accumulator (car battery).

Ans: Lead storage Battery (lead accumulator):

Construction:

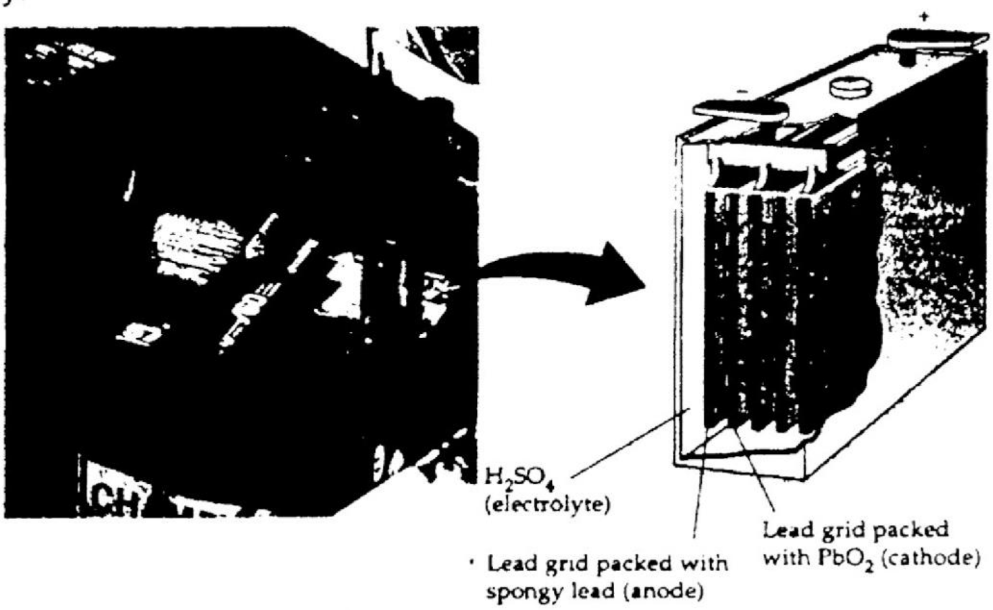
i. In Lead storage Battery anode is a lead plate that becomes coated with $PbSO_4$ as the battery discharges.

ii. The cathode is lead impregnated with PbO_2 which also becomes coated with $PbSO_4$ as the battery discharges.

iii. Both electrodes are immersed in electrolyte which is 30% H_2SO_4 solution having density 1.25g cm^{-3}

Voltage:

The cell produces potential of **two volts**. Automobile batteries **use three or six such cells** joined in series to generate a total electrical potential of **6V or 12V** respectively.



A Lead storage battery

Redox reactions during discharging:

Following reactions take place in it

At anode:

$$Pb(s) + SO_4^{-2}\,(aq) \longrightarrow PbSO_4\,(s) + 2\,\bar{e}$$

At cathode:

$$PbO_2\,(s) + SO_4^{-2}\,(aq) + 4H^+\,(aq) + 2\,\bar{e} \longrightarrow PbSO_4\,(s) + 2H_2O(l)$$

Net reaction during discharge of battery is

$$Pb_{(S)} + PbO_{2(S)} + 2H_2SO_{4(aq)} \longrightarrow 2PbSO_{4\,(S)} + 2H_2O_{\,(l)}$$

Thus during discharge H_2SO_4 is used up and its density decreases. When both the electrodes are completely covered with $PbSO_4$ the battery ceases to deliver current until it is recharged.

Recharging the Battery:

i. Battery can be recharged by connecting its anode to the negative terminal of direct current and the cathode to the positive terminal of the direct current

ii. Reverse chemical reactions occur at anode and cathode of the battery

iii. Thus deposition of Pb on anode and PbO_2 on cathode takes place. In this way we can recharge the battery.

Redox reactions during recharging:

The reactions of recharging of battery are as follows.

At cathode:

$$PbSO_{4(S)} + 2\check{e} \longrightarrow Pb_{(S)} + SO_4^{-2}{}_{(aq)}$$

At anode:

$$PbSO_{4(S)} + 2H_2O_{(l)} \longrightarrow PbO_{2(S)} + SO_4^{-2}{}_{(aq)} + 4H^+_{(aq)} + 2\check{e}$$

Net reaction: $2PbSO_{4(S)} + 2H_2O_{(l)} \longrightarrow Pb_{(S)} + PbO_{2(S)} + 4H^+_{(aq)} + 2SO_4^{-2}{}_{(aq)}$

After recharging H SO solution is concentrated again bringing density to its initial value of 1.25 g cm

Use:

The lead storage battery provides electrical power in automobiles. It is well suited for this use because it supplies the large current needed to drive starter motors and headlights and can be recharged easily.

Q28. What are fuel Cells? How they work? Also write the redox reactions.

Ans: Fuel Cell:

A fuel cell is a special type of galvanic cell in which reactants are continuously supplied as they are consumed and the products are continuously removed.

Principle:

A fuel cell is based upon the reaction between oxygen and a gaseous fuel hydrogen or methane. When hydrogen burns in air, an exothermic reaction occurs. A lot of chemical energy is released in the form of heat and light.

This energy is used for cutting and welding metals. In this reaction hydrogen is oxidized to water.

$$H_2(g) + O_2(g) \xrightarrow{\text{ignite}} H_2O(g) + \text{heat}$$

Working:

i. If the above oxidation-reduction reaction is carried in separate compartments connected in series.

ii. Electrons released in the oxidation of hydrogen in one compartment begin to flow through the external circuit towards the other compartments

iii. There these electrons bring about reduction of oxygen.

iv. Thus electricity begins to flow in the circuit.

v. The energy released from the reaction of hydrogen with oxygen to form water is converted to electrical energy.

Construction:

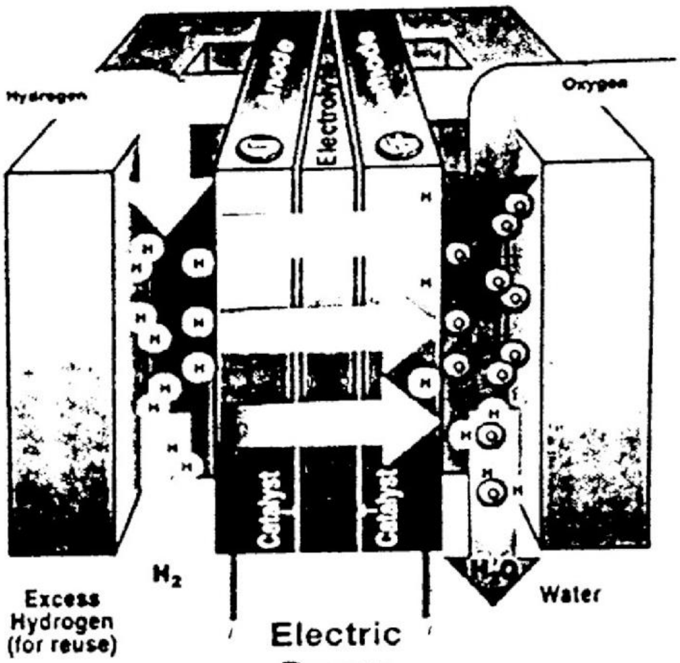
i. A hydrogen-oxygen fuel cell has three compartments separated from one another by porous carbon electrodes.

ii. These electrodes contain platinum as catalyst.

iii. The middle compartment contains a hot aqueous solution of KOH.

iv. Hydrogen gas is passed through the anode compartment and oxygen is passed through the cathode compartment.

Individual Fuel Cell



A Fuel Cell

Redox reaction:

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

Anode reaction = $2H_2(g) + 4OH^-(aq) \longrightarrow 4H_2O(l) + 4\bar{e}$

Cathode reaction = $O_2(g) + 2H_2O(l) + 4\bar{e} \longrightarrow 4OH^-(aq)$

Net cell reaction = $2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$

This is clearly the equation for the burning of hydrogen in oxygen but in a fuel cell it is burning without flame. Thus the electron released in the oxidation of hydrogen flow through the circuit towards the cathode.

- A hydrogen-oxygen cells delivers **0.9 V.**

Note:

The fuel cell operates at high temperature so the water formed evaporates and may be condensed. The water removed in spacecraft is consumed by the astronauts. The fuel cells of this kind have been used by American space program.

Q29. Define solar cell. How can we get electrical energy from it?

Ans: Solar Cell:

Devices that convert solar energy directly into electric energy are called solar cells.

Construction:

i. A semiconductor material is used in these cells.

ii. This material generates voltage output with light input.

iii. A basic solar cell consists of two layers of different types of semi-conductive materials. These materials are joined together to form a junction.

Working:

i. When one layer is exposed to light, many electrons acquire enough energy and break away from their parent atoms. Such electrons cross the junction.

ii. This means that negative ions are formed on one side of the junction and positive ions are on the other side.

iii. Thus a potential difference is developed which causes electrons to flow.

iv. Semiconductor made of silicon gives an output of 0.5 V per cell.

Note:

Research is continuing to get more output with other semi-conductor material. In future solar cells will serve as a cheap source of energy.

Q30. Explain the process of corrosion and how we prevent a metal from it?

Ans: Corrosion:

Corrosion is a natural process, which converts, refined metals to their more stable metal oxides.

Explanation:

The oxidizing agent in corrosion chemistry is atmospheric oxygen. It is most familiar in the form of the rusting of iron. Rusting is an electrochemical process. One of the half reaction in rusting is

$$Fe^{+2} + 2\bar{e} \longrightarrow Fe \qquad E^0 = -0.44V$$

Reverse of this reaction

i.e. $Fe \longrightarrow Fe^{+2} + 2\bar{e}$

is driven by the presence of oxygen. Iron (II) is oxidized further to iron (III) and various insoluble hydrated oxides of iron (III) are deposited as the red-brown precipitate known as rust. These oxides are porous, flake off and expose metal to further corrosion.

The process of corrosion occurs when metal is in contact with water. The water layer present on the surface of iron or a water droplet on its surface dissolves O_2 and CO_2.

In certain industrial areas where SO_2 or other acidic vapours are present also dissolve. Thus metal comes in contact with the electrolyte. The reduction half reaction is:

$$O_2(g) + H^+(aq) + \bar{e} \longrightarrow 2H_2O(l) \qquad E^0 = +1.23V$$

This is more positive than Fe^{+2} - Fe couple, so it can derive the following reaction:

$$Fe \longrightarrow Fe^{+2} + e^- \qquad \text{(oxidation reaction)}$$

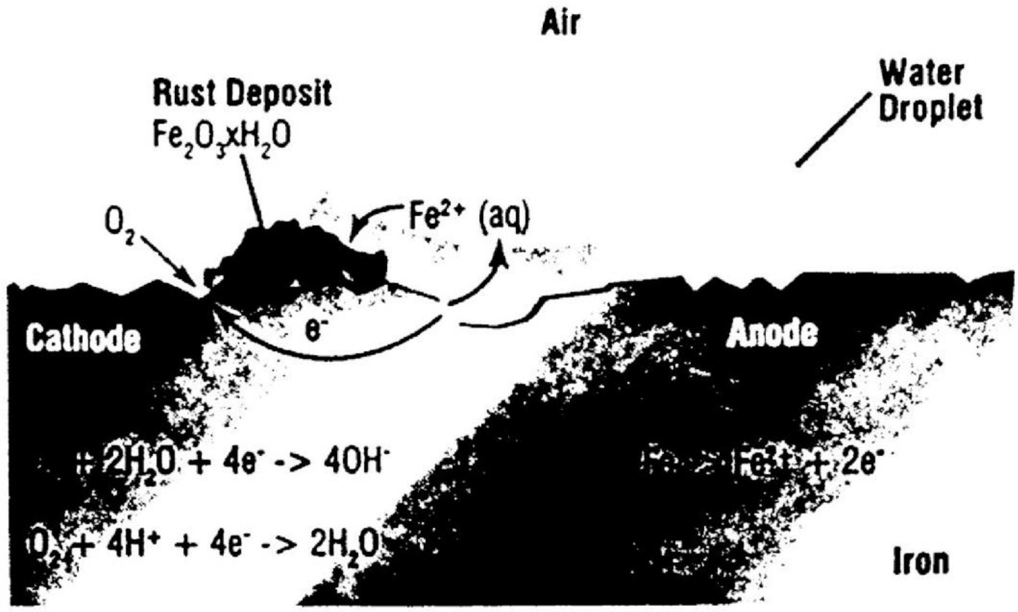
The E^0 cell of the combined half-reactions is

$$E^0\ cell = E^0_{cathode} - E^0_{anode}$$

$$E^0\ cell = +1.23 - (-0.44)$$

$$E^0\ cell = +1.67V$$

Therefore, there is a strong tendency towards oxidation. Oxidation of iron occurs in an interior region of the droplet whereas reduction of O_2 occurs near the air-droplet interface.



Corrosion process of iron

Prevention:

i. Corrosion cannot be eliminated, but sealing the surface from attacks can slow it down.

ii. The corrosion of a metal can be preventing by painting the metal so that it does not come in contact with oxygen and moisture and other harmful agents. Painting also provides visual appeal. That is why bridges, trains, cars etc are painted.

iii. A metal surface can also be protected by coating it with a thin layer of a second metal that is less electropositive than the first.
This can be done by **galvanization or electroplating**.

Q31. How can we protect iron from rusting by using the process of galvanization?

Ans: Galvanization:

Objects made of iron are dipped in molten zinc and dried. This process is known as **galvanization**.

Explanation:

If a scratch penetrates the zinc layer, iron is still protected because Zn oxidizes preferentially. This is because Zn is more active metal than iron, as the potentials for reduction show.

$$Fe^{+2} + 2\bar{e} \rightleftharpoons Fe \quad E^0 = -0.44V$$
$$Zn^{+2} + 2\bar{e} \rightleftharpoons Zn \quad E^0 = -0.76\ V$$

Sacrificial corrosion:

Any oxidation that occurs dissolves Zn rather than Fe. Thus Zn acts as sacrificial coating on Fe. This is also known as sacrificial corrosion.

Q32. Explain the process of electroplating with the help of examples.

Ans: Electroplating:

Electrolysis can be used to deposit one metal on another. A layer of silver or gold is often plated on jewelry and tableware made from inexpensive metals such as iron.

Construction:

i. The article to be plated is used as cathode.

ii. The metal, which is to be deposited on the article, is used as anode.

iii. Water-soluble salt of anode metal is used as electrolyte.

Working:

i. When electrical potential is applied, electrons are ejected from anode and move into the cathode (article).

ii. Metal ions of anode in solution capture the electrons and adhere to the article (cathode).

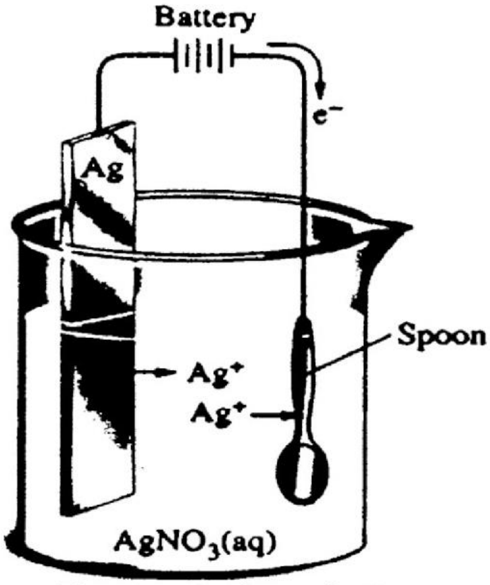
Example:

In silver plating, silver rod is used as anode and sodium cyanide is used as electrolyte. Cynaide ions form a complex with silver ions.

Reactions: $Ag^+ (aq) + 2CN^- (aq) \rightleftharpoons [Ag(CN)_2]^- (aq)$

Anode reaction: $Ag (s) \longrightarrow Ag^+ (aq) + e^-$

Cathode reaction: $[Ag(CN)_2]^- (aq) + e^- \longrightarrow Ag (s) + 2CN^- (aq)$



Electroplating of sliver

Use:

Steel objects are often protected from corrosion by electroplating with Aluminum or Chromium. These metals form a thin protective coating of oxide which inhibits further corrosion. The potential of the passive oxide coating is much like a noble metal.

SUMMARY OF KEY TERMS

1. Redox reactions i.e., oxidation and reduction reactions involve the transfer of electrons or change in oxidation numbers.
2. Redox equations can be balanced by using oxidation number method and ion electron method.
3. The driving force behind the spontaneous redox reaction is called the cell potential.
4. The magnitude of cell potential depends upon the conditions under which the measurement is made. Under standard conditions, all solutions have 1m concentrations; all gases have partial pressure of 1 atm. The standard

potential for the reduction of h^+ to hydrogen gas is arbitrarily taken as zero volts

5. In a galvanic cell oxidation and reduction reactions take place at separate electrodes and electron flow through the external circuit These separate parts of the galvanic cell are half cells. The reactions which occur at these half cells are the half-cell reactions. A salt bridge allows the ions to flow between the half cells.
6. In a galvanic cell, the oxidation occurs at node and reduction occurs at cathode and the electrons flow in the external circuit from anode to cathode.
7. Voltaic cells use a spontaneous redox reaction to drive an electric current through a wire. Whereas, the electrolytic cells use an electric current to drive a redox reaction.
8. The quantity of electricity carried by 1mole of electron is called a faraday It is equal to 96,500 coulombs.
9. In electrolysis electric current from an external source drives a non-spontaneous chemical reaction The amount of chemical reaction that takes place in electrolysis is directly proportional to quantity of charge transferred at the electrode.
10. A battery is a galvanic cell or a group of galvanic cells connected in series. Some of the well-known batteries are the dry cell, the nickel-cadmium battery, lead-storage battery used in automobiles, fuel cells etc.
11. The corrosion of metals is an electrochemical phenomenon.

EXERCISE

MULTIPLE CHOICE QUESTIONS

1: Choose the correct answer.

i. The electrode through which the electron enter the electrolytic solution is

(a) Anode (b) Cathode (c) Salt bridge (d) Electrolyte.

ii. Oxidation number of S in $Na_2S_2O_3$ is

(a) +1 (b) +2 (c) +3 (d) +4

iii. In the electrolysis of molten $ZnCl_2$, the cathode reaction is $Zn^{+2} + 2e^- \rightarrow Zn$ what quantity of electricity is used up for the production of half mole of Zn

(a) 2 Coulombs (b) 1F (c) 2F (d) 2 ampere

iv. How many moles of Cr will be produced by 1.5 Faradays of electricity by the following reaction $Cr^{+3} + 3e^- \longrightarrow Cr$

(a) 0 1 (b) 0.2 (c) 0.03 (d) 1.5 (e) 0.5

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v. **E^0_{Sn} = -0.14V, E^0_{Pb} = -0.13V**

(a) Sn^{+2} can oxidize Pb (b) Pb^{+2} can oxidize Sn

(c) Both can oxidize each other (d) Both can reduce each other.

vi. **A fuel call operates at _______ temperature**

(a) Low(b) Medium (c) Room (d) High

vii. **The oxidation number of Cl in $HClO_4$ is**

(a) –1 (b) +1 (c) +5 (d) +7

viii. **In which of the following compounds oxidation number of N is +5**

(a) NO_2 (b) N_2O_4 (c) N_2O_3 (d) N_2O_5

ix. **The Passage of current through an electrolyte is due to the movement of**

(a) Electrons (b) Anions (c) Cations (d) Ions

x. **Oxidation number of an element in Free State is**

(a) Negative (b) Positive (c) Zero (d) ± 1

xi. **Which of these is not true of an electrolyte.**

(a) It can conduct electricity in molten state

(b) It can conduct electricity in the form of aqueous solution.

(c) It can conduct electricity in the solid form.

(d) It can be an acid, a base or a salt.

xii. **In an electrolytic cell, the name of electrode which has more negative potential begins with the letter**

(a) A to B (b) C to D (c) K to O (d) U to Z

xiii. **Corrosion is an electrochemical process which requires**

(a) Oxygen (b) Water

(c) Acidic vapours (d) Basic vapours

1. a
2. a,b
3. a, b, c
4. a, b, c, d

xiv. **A fuel cell is based upon the reaction between**

(a) Oxygen (b) Gaseous fuel (c) KOH (d) Pt

1. a
2. a,b
3. a, b, c
4. a, b, c, d

Answers

i. b	ii. b	iii. b	iv. e	v. b	vi. d	vii. b
viii. d	ix. d	x. c	xi. c	xii. b	xiii. 3	xiv. 2

2. **Explain, why?**

(a) **Reduction of 1 mole of each Zn^{+2} and Ag^{+} require different Faradays of electricity.**

(b) **It is not possible to measure the potential of an isolated half-cell.**

(c) **The life of a dry cell is shorter than that of an alkaline dry cell.**

Ans: **(a)** One faraday is equal to the charge carried one mole of electrons and it is equal to 96487 C.

One faraday is one mole of electrons.

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

We can see that this requires two moles of electrons per mole of zinc so 2 faradays or about 193000 C charge of electricity is required.

$$Ag^{+}\,(aq) + e^{-} \longrightarrow Ag\,(s)$$

We can see that this requires one mole of electrons per mole of silver so 1 faraday or about 96500 C charge of electricity is required.

(b) A redox reaction always consists of oxidation and reduction. In the cell, the electrode potential develops due to oxidation at anode and reduction, at cathode. Therefore, it is not possible to measure the potential of an isolated half cell.

Example:

If only a zinc half-cell were constructed, no complete redox reaction can occur and so no electrical potential can be measured. It is only measured when another half-cell is combined with the zinc half-cell.

(c) The life of a dry cell is shorter than that of an alkaline dry cell because in alkaline dry cell, a moist past of KOH is used as electrolyte. That is why the alkaline dry cell lasts longer because the zinc anode corrodes less rapidly in basic conditions.

3. **Write the spontaneous reaction for the following sets of half-reactions.**

$Au^{+3}(aq) + 3e^{-} \rightarrow Au(s)$ $\quad E^0 = 1.5\ V$

$NO_3^{-}\,(aq) + 4H^{+}(aq) + 3e^{-} \rightarrow NO(g) + 2H_2O(l)$ $\quad E^0 = 0.096\ V$

Solution:

(i) $Au^{+3}(aq) + 3e^{-} \rightarrow Au(s) \qquad E^\circ = 1.5\ V$

(ii) $NO_3^{-}(aq) + 4H^{+}\,(aq) + 3e^{-} \rightarrow NO(g) + 2H_2O\,(l) \qquad E^\circ = 0.096\ V$

The above equations shows that the reduction potential of first reaction is greater than the second reaction therefore the reduction will occur in the first reaction and oxidation will occur in the second reaction.

Cell reaction:

Cell reaction will occur if we reverse the second reaction and then add it into the first reaction.

At Cathode (Reduction):

$$Au^{+3}(aq) + 3e^{-} \rightarrow Au(s) \qquad E^\circ = 1.5\ V$$

At Anode (Oxidation):

$$NO(g) + 2H_2O\,(l) \rightarrow NO_3^{-}(aq) + 4H^{+}(aq) + 3e^{-} \qquad E^\circ = 0.096\ V$$

Net reaction: $Au^{+3}(aq) + NO(g) + 2H_2O\,(l) \rightarrow NO_3^{-}(aq) + 4H^{+}(aq) + Au(s)$

Cell potential:

$$E^o = E^o_{cathode} - E^o_{anode}$$
$$E^o = (1.5\ V) - (0.96V)$$
$$E^o_{cell} = (1.5\ V) - (0.96V)$$
$$E^o_{cell} = 1.596\ V$$

As the cell potential is positive, therefore, the reaction is spontaneous.

4. Explain the following with reasons.

(a) The oxidation potential of Zn is +0.76V and its reduction potential is –0.76V

(b) A salt bridge maintains the electrical neutrality in the cell.

(c) Na and K can displace hydrogen from acids but Cu and Pt cannot.

(d) Lead storage battery is rechargeable battery.

(e) Zn plating saves Fe from corrosion.

Ans: (a) Oxidation potential of zinc metal is 0.76 V it shows that zinc can lose electron easily and oxidation reaction is feasible. The reduction potential of zinc metal is -0.76 V it shows that zinc cannot reduce easily.

(b) A salt bridge maintains the electrical neutrality in the cell because it only allows the movement of ions from one solution to the other without mixing the two solutions.

(c) It is observed that metal like Na, K can displace H_2 but metals like Cu, Pt cannot displace H_2.

Reason:

i. In active series Na and K are ranked above H_2 therefore they can displace hydrogen.

ii. In active series Cu and Pt are ranked below H_2 therefore they cannot displace hydrogen.

iii. The reduction potential values of Na and K are -2.714V and -2.925V respectively. That is why they can oxidized easily. Hence on reaction with an acid, they displace the Hydrogen easily.

The reduction potential of Cu is +0.521 V and that of Pt is +1.55 V. Due to high values of reduction potentials, Cu is less reactive towards acids and Pt is nearly inert towards acids.

(d) During discharging of lead storage battery, H_2SO_4 is used up, its density decreases. When both the electrodes are completely covered with $PbSO_4$ the battery ceases to deliver current.

Recharging the Battery:

i. Battery can be recharged by connecting its anode to the negative terminal of direct current and the cathode to the positive terminal of the direct current.

ii. Reverse chemical reactions occur at anode and cathode of the battery.

iii. Thus deposition of Pb on anode and PbO_2 on cathode takes place. In this way we can recharge the battery.

Redox reactions during recharging:

The reactions of recharging of battery are as follows.

At cathode:

$$PbSO_{4(S)} + 2\bar{e} \longrightarrow Pb_{(S)} + SO_4^{-2}{}_{(aq)}$$

At anode:

$$PbSO_{4(S)} + 2H_2O_{(l)} \longrightarrow PbO_{2(S)} + SO_4^{-2}{}_{(aq)} + 4H^+_{(aq)} + 2\bar{e}$$

Net reaction: $2PbSO_{4(S)} + 2H_2O_{(l)} \longrightarrow Pb_{(S)} + PbO_{2(S)} + 4H^+_{(aq)} + 2SO_4^{-2}{}_{(aq)}$

After recharging H_2SO_4 solution is concentrated again bringing density to its initial value of 1.25 g cm^{-3}

5. What kind of reaction occurs at the negative terminal of the galvanic cell?

Ans: Oxidation reaction will occur at the negative terminal of the galvanic cell.

Explanation:

In Zn-Cu galvanic cell Zn has greater tendency to lose electrons than Cu. Therefore, Zn electrode acquires negative charge relative to Cu electrode. The electrons flow from Zn electrode through the external circuit to Cu electrode. Thus zinc will act as the negative terminal.

Hence at negative terminal (anode) of cell an oxidation reaction will occur.

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

6. Corrosion is often accelerated where the coating on the body of a car has begun to crack, Explain.

Ans: Body of a car is made up of iron and it is coated with paint. Paint coating save the body of the car from rust. Corrosion is often accelerated where the coating on the body of a car has begun to crack it is because the metal of the car comes in contact with oxygen and moisture and other harmful agents and following reaction would take place.

Reduction half reaction: $O_{2\,(g)} + H^+_{(aq)} + e^- \rightarrow 2H_2O_{(l)}$

Oxidation half reaction: $Fe \rightarrow Fe^{+2} + e^-$ (oxidation reaction)

Conclusion:

Hence oxygen oxidizes Fe to Fe^{+2} which are further oxidized to Fe^{+3} The Fe^{+3} form various insoluble hydrated oxides known as rust and thus the corrosion will accelerated.

7. How many hours would electroplating have to be continued at the rate of 5 amperes if 75g of copper is to be deposited from $CuSO_4$ Solution?

Solution:

Mass of Cu to be produced = 75 g

Atomic mass of Cu = 63.5 g $mole^{-1}$

$$Number\ of\ moles\ of\ Cu = \frac{75g}{63.5g\ mole^-} = 1.1811\ moles$$

Reaction at cathode:

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

The above reaction indicate that for every 2 moles of electrons, 1 mole of Cu will produce.

As we know that

1 F = 1 mole of electron.

2 F = 2 moles of electrons = 1 mole of Cu.

1 mole of Cu is produced by charge = 2 F
1.1811 moles of Cu is produced by charge = 2×1.1811
= 2.3622 F

As we know that **1 F = 96500 C**

2.3622 F = $96500 \text{ C} \times 2.3622$ = 227952.3 C

Time = ?
Charge = 227952.3 C
Current = 5 A

$$Current = \frac{Charge}{Time}$$

$$Time = \frac{Charge}{Current} = \frac{227952.3}{5} = 45590.46\ s$$

$$Time = \frac{45590.46}{3600} = 12.664\ hours$$

8. What is the difference between the following
(a) A galvanic and electrolytic cell
(b) Oxidation half-reaction and reduction half-reaction

Ans: (a) A galvanic and electrolytic cell:

Galvanic Cell	Electrolytic cell
i. A Galvanic cell converts chemical energy into electrical energy.	**i.** An electrolytic cell converts electrical energy into chemical energy.
ii. Here, the redox reaction is spontaneous and is responsible for the production of electrical energy.	**ii.** The redox reaction is not spontaneous and electrical energy has to be supplied to initiate the reaction.
iii. The two half-cells are set up in different containers, being connected through the salt bridge or porous partition.	**iii.** Both the electrodes are placed in a same container in the solution of molten electrolyte.
iv. Here the anode is negative and cathode is the positive electrode. The reaction at the anode is oxidation and that at the cathode is reduction.	**iv.** Here, the anode is positive and cathode is the negative electrode. The reaction at the anode is oxidation and that at the cathode is reduction.
v. The electrons are supplied by the species getting oxidized. They move from anode to the cathode in the external circuit.	**v.** The external battery supplies the electrons. They enter through the cathode and come out through the anode.

(b) Oxidation half-reaction and reduction half-reaction:

Oxidation half-reaction	Reduction half-reaction
i. Reaction taking place in oxidation half-cell is called oxidation half reaction.	**i.** The reaction taking place in reduction half-cell is called reduction half reaction.
ii. In oxidation half-reaction, the species is oxidized by losing electrons.	**ii.** In reduction half-reaction, the species is reduced by gaining electrons.

iii. This process occurs at anode. iii. This process occurs at cathode

iv. Oxidation half reaction. iv. Reduction half reaction:

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

9. Sketch the galvanic cell based on the following overall reaction. Show direction of electron flow and identify the anode, write balanced reaction. Assume all concentrations are 1.0 M

(a) $Cr^{+3}(aq) + Cl_2(aq) \longrightarrow Cr_2O_7^{\,2}(aq) + Cl^-(aq)$

(b) $Cu^{+2}(aq) + Mg(s) \longrightarrow Mg^{+2}(aq) + Cu(s)$

Ans: (a) $Cr^{+3}_{(aq)} + Cl_{2(aq)}$ $Cr_2O_7^{-2}{}_{(aq)} + Cl_{(aq)}$

i. In this reaction Cl atom changes its oxidation state from zero in Cl_2 to -1 in Cl^-. It is a reduction process as it gain $1e^-$ per Cl atom.

ii. In this reaction the Cr atom changes its oxidation state from +3 in Cr^{+3} to +6 in $Cr_2O_7^{-2}$. It is a oxidation process as it loss $3e^-$ per Cr atom

At Anode: (Oxidation half reaction)

$$2Cr^{+3}_{(aq)} + 7H_2O \rightarrow Cr_2O_7^{-2}{}_{(aq)} + 14H^+ + 6e^- \quad ....(i)$$

At Cathode: (Reduction half reaction)

$$Cl_2 + 2e \rightarrow 2Cl^{-1}{}_{(aq)} \quad ....(ii)$$

Electron flow:

Thus the electron flows from **anode to cathode.**

Balancing:

To balance loss and gain of electrons

multiply Eq. (ii) by 3 and then add to Eq. (i) to get overall cell reaction.

Oxidation half reaction: $2Cr^{+3}_{(aq)} + 7H_2O \rightarrow Cr_2O_7^{-2}{}_{(aq)} + 14H^+ + 6e^-$

Reduction half reaction: $3Cl_{2(aq)} + 6e^- \rightarrow 6Cl^{-1}{}_{(aq)}$

Overall Cell Reaction. $2Cr^{+3}_{(aq)} + 7H_2O + 3Cl_{2(aq)} \rightarrow Cr_2O_7^{-2}{}_{(aq)} + 14H^+ + 6Cl^{-1}{}_{(aq)}$

(b) $Cu^{+2}(aq) + Mg(s) \longrightarrow Mg^{+2}(aq) + Cu(s)$

i. The Cu atom changes its oxidation state from +2 in Cu^{+2} to 0 in Cu. It is a reduction process as it gain $2e^-$ per atom

ii. In this reaction the Mg atom changes its oxidation state from 0 in Mg to +2 in Mg^{+2}. It is a oxidation process as it gain $2e^-$ per atom.

At Anode: (Oxidation half reaction)

$$Mg(s) \longrightarrow Mg^{+2} + 2e^- \quad ...(i)$$

At Cathode: (Reduction half reaction)

$$Cu_2(aq) + 2e^- \longrightarrow Cu_{(aq)} \quad ....(ii)$$

Electron flow:

Thus the electron flows from **anode to cathode.**

Balancing:

Write equation (i) and (ii) as it is

Oxidation half reaction: $Mg(s) \longrightarrow Mg^{+2} + 2e^-$

Reduction half reaction: $Cu_2(aq) + 2e^- \longrightarrow Cu_{(aq)}$

Net Reaction: $Cu^{+2}(aq) + Mg(s) \longrightarrow Mg^{+2}(aq) + Cu(s)$

Q.10: Calculate E^0 for the following cells, which reactions are spontaneous as written under standard conditions.

(a) $2Ag^+(aq) + Cu(s) \longrightarrow Cu^{+2}(aq) + 2Ag(s)$

(b) $Zn^{+2}(aq) + Ni(s) \longrightarrow Ni^{+2}(aq) + Zn(s)$

(c) $Al(s) + 3H^+(aq) \longrightarrow Al^{+3}(aq) + \frac{3}{2}H_2(g)$

Solution: (a) $2Ag^+_{(aq)} + Cu_{(s)} \rightarrow Cu^{+2}_{(aq)} + 2Ag_{(s)}$

Reduction half reaction:

Ag atom changes its oxidation state from +1 in Ag^+ to zero in Ag. It is because of the gain of $1e^-$ Thus it is a reduction process.

At Cathode: (Reduction half reaction)

$$2Ag^+_{(aq)} + 2e^- \rightarrow 2Ag_{(aq)}$$

Oxidation half reaction:

Cu atom changes its oxidation state from zero in Cu to +2 in Cu^{+2} It is because of the loss of $2e^-$ Thus it is an oxidation process.

At Anode: (Oxidation half reaction)

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

Cell potential:

The standard reduction potentials of Cu and Ag are

E^o_{Cu} = +0.34V and E^o_{Ag} = +0.80 V

$E^o_{cell} = E^o_{cathode} - E^o_{anode}$

$E^o_{cell} = E^o_{Ag} - E^o_{Cu}$

$E^o_{cell} = (0.80) - (0.34)$

$E^o_{cell} = 0.46V$

(b) $Zn^{+2}_{(aq)} + Ni_{(s)} \rightarrow Ni^{+2}_{(aq)} + Zn_{(s)}$

Reduction half reaction:

Zn atom changes its oxidation state from +2 in Zn^{+2} to zero in Zn. It is because of the gain of $2e^-$ Therefore it is a reduction process.

At Cathode: (Reduction half reaction)

$$Zn^{+2}_{(aq)} + 2e^- \rightarrow Zn_{(aq)}$$

Oxidation half reaction:

Ni atom changes its oxidation state from zero in Ni to +2 in Ni^{+2} It is because of the loss of $2e^-$ Therefore it is an oxidation process.

The balanced reaction at anode and cathode would be

At Anode: (Oxidation half reaction)

$$Ni_{(s)} \rightarrow Ni^{+2}_{(aq)} + 2e^-$$

Cell potential:

The standard reduction potentials of Ni and Zn are

E^o_{Ni} = -0.25 V and E°_{Zn} = -0.76 V

$E^o_{cell} = E^o_{cathode} - E^o_{anode}$

$E^o_{cell} = E^o_{Zn} - E^o_{Ni}$

$E^o_{cell} = (-0.76) - (-0.25)$

E^o_{cell} = - 0.51V

As the cell potential is negative, thus, this reaction is not spontaneous (not feasible)

(c) $Al_{(s)} + 3H^+_{(aq)} \rightarrow Al^{+3}_{(aq)} + \frac{3}{2} H_{2(g)}$

Oxidation half reaction:

In this reaction, the Al atom changes its oxidation state from zero in Al to +3 in Al^{+3} It is because of the loss of $3e^-$ Thus it is an oxidation process.

Anode: (Oxidation half reaction)

$$Al_{(s)} \rightarrow Al^{+3}_{(aq)} + 3e^-$$

Reduction half reaction:

The H atom changes its oxidation state from +1 in H^+ to zero in H_2. This is due to gain of $1e^-$ Thus it is a reduction process.

At Cathode: (Reduction half reaction)

$$3H^+_{(aq)} + 3e^- \rightarrow \frac{3}{2}H_{2(g)}$$

Cell potential:

Standard reduction potentials of hydrogen and Al are

$E°_{Al}$ = -1.67 V and $E°_H$ = 0.00 V

$E°_{cell} = E°_{cathode} - E°_{anode}$

$E°_{cell} = E°_H - E°_{Al}$

$E°_{cell} = (0.00) - (-1.67)$

$E°_{cell} = +1.67$ V

As the cell potential is positive, thus, this reaction is spontaneous (feasible)

Q.11: An electroplating apparatus in used to coat jewelry with gold. What mass of gold can be deposited from a solution that contains $\{Au(CN)_4\}^{-1}$ ion if a current of 5.0 amperes flows for 30 minutes, following half-reaction occurs.

$$\{Au(CN)_4\}^{-1}(aq) + 3e^- \longrightarrow Au(s) + 4CN^-(aq)$$

Solution: **Charge calculation:**

Time = 30 minutes = 30 × 60 = 1800 s

Current = 5 A

As

$$Charge = \frac{Charge}{Time\ (s)}$$

Charge = Current × Time (s) = 5 A × 1800 s = 9000 C

As

96500 C = 1F

9000 C $= \frac{1F}{96500C} \times 9000C = 9.326 \times 10^{-2}F$

Reaction at cathode :

$$\{Au(CN)_4\}_{(aq)} + 3e^- \rightarrow Au_{(s)} + 4CN^-_{(aq)}$$

It indicates that every 3 moles of electrons, 1 mole of Au will produce.

1 F = 1 mole of electron.

3 F = 3 moles of electrons = 1 mole of Au.

So,

3 F charge produces Au = 1 mole

9.326×10^{-2} F charge produces Au $= \frac{1}{3} \times 9.326 \times 10^{-2}$

$= 3.109 \times 10^{-2}$ moles

Atomic mass of Au $= 179\ g\ mole^{-1}$

1 mole of Au $= 179\ g$

3.109×10^{-2} moles $= 179 \times 3.109 \times 10^{-2} = 5.565\ g$

Q.12: Determine oxidation number of all the atoms in the following

(a) K_2CO_3 (b) P_4 (c) XeF_4 (d) F_2O

(e) PCl_5 (f) NH_3 (g) Fe_3O_4

Ans:

(a) K_2CO_3

Oxidation number of K $= +1$

Oxidation number of O $= -2$

Oxidation number of C $= x$

K_2CO_3

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

$x - 4 = 0$

$x = +4$

(b) P_4

In free state oxidation state of an element is zero. Thus oxidation state of P_4 is zero.

(c) XeF_4

Oxidation number of F $= -1$

Oxidation number of Xe $= x$

XeF_4

$x + 4(-1) = o$

$x - 4 = 0$

$x = +4$

(d) F_2O

oxidation number of F $= -1$

Oxidation number of O $= x$

F_2O

$2(-1) + x = 0$

$-2 + x = 0$

$x = +2$

(e) PCl_5

Oxidation number of Cl $= -1$

Oxidation number of P $= x$

PCl_5

$x + 5(-1) = 0$

$x - 5 = 0$

$x = +5$

(f) NH_3

Oxidation number of H $= +1$

Oxidation number of N $= x$

NH_3

$x + 3(+1) = 0$

$x + 3 = 0$

$x = -3$

(g) Fe_3O_4

Oxidation number of O $= -2$

Oxidation number of Fe $= x$

$$Fe_3O_4$$

$$3x + 4(-2) = 0$$

$$3x - 8 = 0$$

$$3x = +8$$

$$x = +\frac{8}{3}$$

Q.13: Balance the following half-reactions

(a) $Cu^{+} \longrightarrow CuO$ **(b) $I \longrightarrow IO_3^{-1}$**

(c) $S \longrightarrow H_2S$ **(d) $SO_4^{-2} \longrightarrow SO_2$**

Ans: (a) $Cu^{+} \rightarrow CuO$

(i) Balancing of oxygen:

There is one O atom on the right and ,none on the left. So we will add one H_2O on the L.H.S. to balance O atoms.

$$Cu^{+} + H_2O \rightarrow CuO$$

(ii) Balancing of hydrogen:

There are 2 H atoms on the left and none on the right, so we will add $2H^{+}$ on the R H.S. to balance H.

$$Cu^{+} + H_2O \rightarrow CuO + 2H^{+}$$

(iii) Balance Charge:

Now. balance charges the left side has mono-positive charge corresponding to +1

The right side has 2 mono-positive charges corresponding to +2. Thus, right side needs $1e^{-}$

$$Cu^{+} + H_2O \rightarrow CuO + 2H^{+} + 1e^{-}$$

(b) $I \rightarrow IO_3^{-1}$

(i) Balancing of oxygen:

There are 3 O atoms on the right and none on the left. So we will add 3 H_2O on the L.H.S. to balance O atoms.

$$I + 3H_2O \rightarrow IO_3^{-1}$$

(ii) Balancing of hydrogen:

There are 6 H atoms on the left and none on the right, so we will add $6H^{+}$ on the R.H.S. to balance H.

$$I + 3H_2O \rightarrow IO_3^{-1} + 6H^{+}$$

(iii) Balance Charges:

Now, balance charges the left side has no charge corresponding to 0.

The right side has 6 mono-positive and mono-negative charge corresponding to $-1 + 6(+1) = +5$. Thus, right side needs $5e^{-}$

$$I + 3H_2O \rightarrow IO_3^{-1} + 6H^{+} + 5e^{-}$$

(c) $S \rightarrow H_2S$

(i) There are 2 hydrogen atoms on the right side and none on the left side so add $2H^{+}$ on L.H.S. to balance hydrogen.

$$S + 2H^{+} \rightarrow H_2S$$

(ii) Balance Charges:

Now balance charges the left side has 2 mono-positive charge correspond to $2(+1) = +2$. No charges present on right hand side thus left hand side needs $2e$

$$S + 2H^{+} + 2e^{-} \rightarrow H_2S$$

(d) $SO_4^{-2} \rightarrow SO_2$

(i) Balancing of oxygen:

There are 4 oxygen atoms on the left and 2 oxygen atoms on the right to balance oxygen add $2H_2O$ on the right side.

$$SO_4^{-2} \rightarrow SO_2 + 2H_2O$$

(ii) Balancing of hydrogen:

There are 4 hydrogen atoms on the right and none on the left to balance hydrogen add $2H^{+}$ on L.H.S.

$$SO_4^{-2} + 4H^{+} \rightarrow SO_2 + 2H_2O$$

(iii) Balance Charge:

There are 4 mono-positive and 1 di-negative charge on the left side correspond to $(-2) + 4(+1) = +2$.

Right side has no charge thus left side $2e^{-}$

$$SO_4^{-2} + 4H^{+} + 2e^{-} \rightarrow SO_2 + 2H_2O$$

Q.14: Balance the following reactions by ion electron method

(a) $SO_2 + MnO_4^{-} \longrightarrow Mn^{+2} + HSO_4^{-}$

(b) $BrO_3^{-2} + SO_2 \longrightarrow HSO_4^{-} + Br^{-}$

(c) $2N_2O_4 \longrightarrow NO_3^{-} + NO$

(d) $Sn^{+2} + Fe^{+3} \longrightarrow Sn^{+4} + Fe^{+2}$

(e) $H_2O_2 + MnO_4^{-} \longrightarrow Mn^{+2} + O_2$

(f) $Cr_2O_7^{-2} + Cl^{-1} \longrightarrow Cr^{+3} + Cl_2$

(g) $IO_3^{-1} + AsO_3^{-3} \longrightarrow I^{-1} + AsO_4^{-3}$

(h) $H_3AsO_3 + Cr_2O_7^{-2} \longrightarrow H_3AsO_4 + Cr^{+3}$

(i) $Cu + NO_3^{-} \longrightarrow Cu^{+2} + NO$

Solution:

(a) $SO_2 + MnO_4^{-} \longrightarrow Mn^{+2} + HSO_4^{-}$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

$$MnO_4^{-1} \rightarrow Mn^{+2}$$

Oxidation half reaction:

$$SO_2 \rightarrow HSO_4^{-1}$$

Step 2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

Mn atoms are already balanced.

$$MnO_4^{-1} \rightarrow Mn^{+2}$$

Add $4H_2O$ on R.H.S. to balance O atoms.

$$MnO_4^{-1} \rightarrow Mn^{+2} + 4H_2O$$

Add $8H^{+}$ on L.H.S. to balance H.

$$MnO_4^{-1} + 8H^{+} \rightarrow Mn^{+2} + 4H_2O$$

(ii) Balance Charge:

The total charge on L.H.S. is -1+8(+1) = + 7 The total charge on R.H.S. is +2. Thus, add + $5e^-$ on L.H.S.

$MnO_4^{-1} + 8H^+ + 5e^- \rightarrow Mn^{+2} + 4H_2O$ ....(1)

Oxidation half reaction:

(i) Balance Mass:

S atoms are already balanced.

$SO_2 \rightarrow HSO_4^{-1}$

Add $4H_2O$ on L.H.S. to balance O atoms.

$SO_2 + 4H_2O \rightarrow HSO_4^{-1}$

Add $3H^+$ on R.H.S. to balance H atoms.

$SO_2 + 2H_2O \rightarrow HSO_4^{-1} + 3H^+$

(ii) Balance Charge:

The L.H.S. has no charge. The total charge on R.H.S. is (-1)+3(+1) =+2. Thus, add $2e^-$ on R.H.S.

$SO_2 + 2H_2O \rightarrow HSO_4^{-1} + 3H^+ + 2e^-$ ....(2)

Step 3. Balance the loss and gain of electrons in both half reactions and add them.

The reduction half-reaction has $5e^-$ while the oxidation half-reaction has $2e^-$ Thus multiply Eq. (1) by 2, Eq. (2) by 5 and then add the resulting equations.

$$2MnO_4^{-1} + 16H^+ + 10e^- \rightarrow 2Mn^{+2} + 8H_2O$$

$$5SO_2 + 10H_2O \rightarrow 5HSO_4^{-1} + 15H^+ + 10e^-$$

$$2MnO_4^{-1} + 16H^+ + 10e^- + 5SO_2 + 10H_2O \rightarrow 2Mn^{+2} + 8H_2O + 5HSO_4^{-1} + 15H^+ + 10e^-$$

Step 4: Strike out duplication of species on both the sides.

Duplications are $10e^-$, $15H^+$ and $8H_2O$. Cancel this out from both the sides.

$$2MnO_4^{-1} + H^+ + 5SO_2 + 2H_2O \rightarrow 2Mn^{+2} + 5HSO_4^{-1}$$

(b) $BrO_3^{-2} + SO_2 \rightarrow HSO_4^{-1} + Br^-$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

$BrO_3^{-2} \rightarrow Br^{-1}$

Oxidation half reaction:

$SO_2 \rightarrow HSO_4^{-1}$

Step 2: Balance each half reaction separately.

Reduction half reaction:

(ii) Balance Mass:

Br atoms are already balanced.

$BrO_3^{-2} \rightarrow Br^{-1}$

Add $3H_2O$ on R.H.S. to balance O atoms.

$BrO_3^{-2} \rightarrow Br^{-1} + 3H_2O$

Add $6H^+$ on L.H.S. to balance H atoms.

$BrO_3^{-2} + 6H^+ \rightarrow Br^{-1} + 3H_2O$

(ii) Balance Charge:

The total charge on L.H.S. is -2+6(+1) = +4. The total charge on R.H.S is -1. Thus, add $5e^-$ on L.H.S.

$$BrO_3^{-2} + 6H^+ + 5e \rightarrow Br^{-1} + 3H_2O \quad (1)$$

Oxidation half reaction:

(i) Balance Mass:

S atoms are already balanced

$$SO_2 \rightarrow HSO_4^{-1}$$

Add $3H_2O$ on L.H.S. to balance O atoms

$$SO_2 + 4H_2O \rightarrow HSO_4^{-1}$$

Add $3H^+$ on R.H.S. to balance H atoms.

$$SO_2 + 2H_2O \rightarrow HSO_4^{-1} + 3H^+$$

(ii) Balance Charge:

The L.H.S. has no charge. The total charge on R.H.S. is (-1)+3(+1)=+2. Thus, add $2e^-$ on R.H.S.

$$SO_2 + 2H_2O \rightarrow HSO_4^{-1} + 3H^+ + 2e^- \quad (2)$$

Step 3: Balance the loss and gain of electrons in both half reactions and add them.

The reduction half-reaction has $5e^-$ while the oxidation half-reaction has $2e^-$ Thus multiply Eq. (1) by 2, Eq. (2) by 5 and then add the resulting equations.

$$2BrO_3^{-2} + 12H^+ + 10e^- \rightarrow 2Br^{-1} + 6H_2O$$

$$5SO_2 + 10H_2O \rightarrow 5HSO_4^{-1} + 15H^+ + 10e^-$$

$$2BrO_3^{-2} + 12H^+ + 10e^- + 5SO_2 + 10H_2O \rightarrow 2Br^{-1} + 6H_2O + 5HSO_4^{-1} + 15H^+ + 10e^-$$

Step 4: Strike out duplication of species on both the sides.

Duplications are $10e^-$, $12H^+$ and $6H_2O$. Cancel this out from both the sides.

$$2BrO_3^{-2} + 5SO_2 + 4H_2O \rightarrow 2Br^{-1} + 5HSO_4^{-1} + 3H^+$$

(c) $2N_2O_4 \rightarrow NO_3^{-1} + NO$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

$$N_2O_4 \rightarrow NO$$

Oxidation half reaction:

$$N_2O_4 \rightarrow NO_3^{-1}$$

Step2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

Use 2 before NO on R.H.S. to balance N atoms.

$$N_2O_4 \rightarrow 2NO$$

Add $2H_2O$ on R.H.S. to balance O atoms.

$$N_2O_4 \rightarrow 2NO + 2H_2O$$

Add $4H^+$ on L.H.S. to balance H.

$$N_2O_4 + 4H^+ \rightarrow 2NO + 2H_2O$$

(ii) Balance Charge:

The total charge on L.H.S. is 4(+l) = +4. The R.H.S. has no charge. Thus, add $4e^-$ on L.H.S.

$$N_2O_4 + 4H^+ + 4e \rightarrow 2NO + 2H_2O \quad (1)$$

Oxidation half reaction:

(i) Balance Mass:

Use 2 before NO_3^{-1} on R.H.S. to balance N atoms.

$$N_2O_4 \rightarrow 2NO_3^{-1}$$

Add $2H_2O$ on L.H.S. to balance O atoms.

$$N_2O_4 + 2H_2O \rightarrow 2NO_3^{-1}$$

Add $4H^+$ on R.H.S. to balance H atoms

$$N_2O_4 + 2H_2O \rightarrow 2NO_3^{-1} + 4H^+$$

(ii) Balance Charge:

The L.H.S. has no charge. The total charge on R.H.S. is 2(-1)+4(+1) = +2. Thus, add $2e^-$ on R.H.S.

$$N_2O_4 + 2H_2O \rightarrow 2NO_3^{-1} + 4H^+ + 2e \ldots\ldots\ldots\ldots\ldots\ldots (2)$$

Step3: Balance the loss and gain of electrons in both half reactions and add them.

The reduction half-reaction has $4e^-$ while the oxidation half-reaction has $2e^-$ Thus multiply Eq (2) by 2 and then add to Eq. (1).

$$N_2O_4 + 4H^+ + 4e^- \rightarrow 2NO + 2H_2O$$
$$2N_2O_4 + 4H_2O \rightarrow 4NO_3^{-1} + 8H^+ + 4e^-$$

$$3N_2O_4 + 4H^+ + 4H_2O + 4e^- \rightarrow 2NO + 2H_2O + 4NO_3^{-1} + 8H^+ + 4e^-$$

Step4: Strike out duplication of species on both the sides.

Duplications are $4e^-$, $4H^+$ and $2H_2O$. Cancel this out from both the sides.

$$3N_2O_4 + 2H_2O \rightarrow 2NO + 4NO_3^{-1} + 4H^+$$

(d) $Sn^{+2} + Fe^{+3} \rightarrow Sn^{+4} + Fe^{+2}$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

$$Fe^{+3} \rightarrow Fe^{+2}$$

Oxidation half reaction:

$$Sn^{+2} \rightarrow Sn^{+4}$$

Step2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

Fe atoms are already balanced.

$$Fe^{+3} \rightarrow Fe^{+2}$$

(ii) Balance Charge:

The total charge on L.H.S. is +3. The total charge on R.H.S. is +2. Thus, add $4e^-$ on L.H.S.

$$Fe^{+3} + 1e^- \rightarrow Fe^{+2} \ldots\ldots\ldots\ldots\ldots\ldots (1)$$

Oxidation half reaction:

(i) Balance Mass:

Sn atoms are already balanced.

$$Sn^{+2} \rightarrow Sn^{+4}$$

(ii) Balance Charge:

The total charge on L.H.S. is +2. The total charge on R.H.S. is +4. Thus, add $2e^-$ on R.H.S.

$$Sn^{+2} \rightarrow Sn^{+4} + 2e^- \text{.......................} (2)$$

Step 3: Balance the loss and gain of electrons in both half reactions and add them.

The reduction half-reaction has $1e^-$ while the oxidation half-reaction has $2e^-$ Thus multiply Eq. (1) by 2 and then add to Eq. (2).

$$2Fe^{+3} + 2e^- \rightarrow 2Fe^{+2}$$
$$Sn^{+2} \rightarrow Sn^{+4} + 2e^-$$

$$2Fe^{+3} + 2e^- + Sn^{+2} \rightarrow 2Fe^{+2} + Sn^{+4} + 2e^-$$

Step 4: Strike out duplication of species on both the sides.

Duplications are $4e^-$, $4H^+$ and $2H_2O$. Cancel this out from both the sides.

$$2Fe^{+3} + Sn^{+2} \rightarrow 2Fe^{+2} + Sn^{+4}$$

(e) $H_2O_2 + MnO_4^{-1} \rightarrow Mn^{+2} + O_2$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

$$MnO_4^{-1} \rightarrow Mn^{+2}$$

Oxidation half reaction:

$$H_2O_2 \rightarrow O_2$$

Step2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

Mn atoms are already balanced.

$$MnO_4^{-1} \rightarrow Mn^{+2}$$

Add $4H_2O$ on R.H.S. to balance O atoms

$$MnO_4^{-1} \rightarrow Mn^{+2} + 4H_2O$$

Add $8H^+$ on L.H.S. to balance H

$$MnO_4^{-1} + 8H^+ \rightarrow Mn^{+2} + 4H_2O$$

(ii) Balance Charge:

The total charge on L.H.S. is -1+8(+1) = +7. The total charge on R.H.S. is +2. Thus, add $5e^-$ on L.H.S.

$$MnO_4^{-1} + 8H^+ + 5e^- \rightarrow Mn^{+2} + 4H_2O \text{.......................} (1)$$

Oxidation half reaction:

(i) Balance Mass:

O atoms are already balanced.

$$H_2O_2 \rightarrow O_2$$

Add $2H^+$ on R.H.S. to balance H atoms.

$$H_2O_2 \rightarrow O_2 + 2H^+$$

(ii) Balance Charge:

The L.H.S. has no charge. The total charge on R.H.S. is 2(+1) = +2. Thus, add $2e^-$ on R.H.S.

$$H_2O_2 \rightarrow O_2 + 2H^+ + 2e^- \text{.......................} (2)$$

Step3: Balance the loss and gain of electrons in both half reactions and add them.

The reduction half-reaction has $5e^-$ while the oxidation half-reaction has $2e^-$ Thus multiply Eq. (1) by 2 by 5 and then add the resulting equations

$$2MnO_4^{-1} + 16H^+ + 10e^- \rightarrow 2Mn^{+2} + 8H_2O$$

$$5H_2O_2 \rightarrow 5O_2 + 10H^+ + 10e^-$$

$$2MnO_4^{-1} + 16H^+ + 10e^- + 5H_2O_2 \rightarrow 2Mn^{+2} + 8H_2O + 5O_2 + 10H^+ + 10e^-$$

Step 4: Strike out duplication of species on both the sides.

Duplications are $10e^-$ and $10H^+$. Cancel these out from both the sides.

$$2MnO_4^{-1} + 6H^+ + 5H_2O_2 \rightarrow 2Mn^{+2} + 8H_2O + 5O_2$$

(f) $Cr_2O_7^{-2} + Cl^{-1} \rightarrow Cr^{+3} + Cl_2$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

$$Cr_2O_7^{-2} \rightarrow Cr^{+3}$$

Oxidation half reaction:

$$Cl^{-1} \rightarrow Cl_2$$

Step2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

Use 2 before Cr^{+3} on R.H.S. to balance Cr atoms.

$$Cr_2O_7^{-2} \rightarrow 2Cr^{+3}$$

Add $7H_2O$ on R.H.S. to balance O atoms.

$$Cr_2O_7^{-2} \rightarrow 2Cr^{+3} + 7H_2O$$

Add $14H^+$ on L.H.S. to balance H atoms

$$Cr_2O_7^{-2} + 14H^+ \rightarrow 2Cr^{+3} + 7H_2O$$

(ii) Balance Charge:

The total charge on L.H.S. is -2+14(+1) = +12. The total charge on R.H.S. is 2(+3) = +6. Thus, add $6e^-$ on L.H.S.

$$Cr_2O_7^{-2} + 14H^+ + 6e^- \rightarrow 2Cr^{+3} + 7H_2O \text{ ....(1)}$$

Oxidation half reaction:

(i) Balance Mass:

Use 2 before Cl^{-1} ions on R.H.S. to balance Cl atoms.

$$2Cl^{-1} \rightarrow Cl_2$$

(ii) Balance Charge:

The total charge on L.H.S. is 2(-1) = -2. The R.H.S. has no charge. Thus, add $2e^-$ on R.H.S.

$$2Cl^{-1} \rightarrow Cl_2 + 2e^- \text{......................(1)}$$

Step3: Balance the loss and gain of electrons in both half reactions and add them.

The reduction half-reaction has $6e^-$ while the oxidation half-reaction has $2e^-$ Thus multiply Eq. (2) by 3 and then add to Eq. (1).

$$Cr_2O_7^{-2} + 14H^+ + 6e^- \rightarrow 2Cr^{+3} + 7H_2O$$
$$6Cl^{-1} \rightarrow 3Cl_2 + 2e^-$$

$$Cr_2O_7^{-2} + 14H^+ + 6Cl^{-1} + 6e^- \rightarrow 2Cr^{+3} + 7H_2O + 3Cl_2 + 2e^-$$

Step 4: Strike out duplication of species on both the sides.

Duplications are $6e^-$ Cancel this out from both the sides.

$$Cr_2O_7^{-2} + 14H^+ + 3Cl_2 + 6Cl^{-1} \rightarrow 2Cr^{+3} + 7H_2O + 3Cl_2$$

(g) $IO_3^{-1} + AsO_3^{-3} \rightarrow I^{-1} + AsO_4^{-3}$

Step1: Split the reaction into two half reactions.

Reduction half reaction: $IO_3^{-1} \rightarrow I^{-1}$

Oxidation half reaction: $AsO_3^{-3} \rightarrow AsO_4^{-3}$

Step2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

I atoms are already balanced.

$$IO_3^{-1} \rightarrow I^{-1}$$

Add $3H_2O$ on R.H.S. to balance O atoms.

$$IO_3^{-1} \rightarrow I^{-1} + 3H_2O$$

Add $6H^+$ on R.H.S. to balance H atoms.

$$IO_3^{-1} + 6H^+ \rightarrow I^{-1} + 3H_2O$$

(ii) Balance Charge:

The total charge on L.H.S. is -1+6(+1) = +5. The total charge on R.H.S. is -1. Thus, add $5e^-$ on L.H.S.

$$IO_3^{-1} + 6H^+ + 6e^- \rightarrow I^{-1} + 3H_2O \text{ ....(1)}$$

Oxidation half reaction:

(i) Balance Mass:

S atoms are already balanced.

$$AsO_3^{-3} \rightarrow AsO_4^{-3}$$

Add H_2O on L.H.S. to balance O atoms.

$$AsO_3^{-3} + H_2O \rightarrow AsO_4^{-3}$$

Add $2H^+$ on L.H.S. to balance H atoms

$$AsO_3^{-3} + H_2O \rightarrow AsO_4^{-3} + 2H^+$$

(ii) Balance Charge:

The total charge on L.H.S. is -3. The total charge on R.H.S. is -3+2(+1) = -1. Thus, add $2e^-$ on R.H.S.

$$AsO_3^{-3} + H_2O \rightarrow AsO_4^{-3} + 2H^+ + 2e^- \text{ ....(2)}$$

Step3: Balance the loss and gain of electrons in both half reactions and add them.

The reduction half-reaction has $6e^-$ while the oxidation half-reaction has $2e^-$ Thus multiply Eq. (2) by 3 and then add to Eq. (1).

$$IO_3^{-1} + 6H^+ + 6e^- \rightarrow I^{-1} + 3H_2O$$
$$3AsO_3^{-3} + 3H_2O \rightarrow 3AsO_4^{-3} + 6H^+ + 6e^-$$

$$IO_3^{-1} + 6H^+ + 6e^- + 3AsO_3^{-3} + 3H_2O \rightarrow I^{-1} + 3H_2O + 3AsO_4^{-3} + 6H^+ + 6e^-$$

Step 4: Strike out duplication of species on both the sides.

Duplications are $6e^-$, $6H^+$ and $3H_2O$. Cancel this out from both the sides.

$$IO_3^{-1} + 3AsO_3^{-3} \rightarrow I^{-1} + 3AsO_4^{-3}$$

(h) $H_3AsO_3 + Cr_2O_7^{-2} \rightarrow H_3AsO_4 + Cr^{+3}$

Step1: Split the reaction into two half reactions.

Reduction half reaction:

$$Cr_2O_7^{-2} + \rightarrow Cr^{+3}$$

Oxidation half reaction

$$H_3AsO_3 \rightarrow H_3AsO_4$$

Step2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

Use 2 before Cr^{+3} on R.H.S. to balance Cr atoms.

$$Cr_2O_7^{-2} \rightarrow Cr^{+3}$$

Add $7H_2O$ on R.H.S. to balance O atoms.

$$Cr_2O_7^{-2} \rightarrow Cr^{+3} + 7H_2O$$

Add $14H^+$ on L.H.S. to balance H atoms

$$Cr_2O_7^{-2} + 14H^+ \rightarrow 2Cr^{+3} + 7H_2O$$

(ii) Balance Charge:

The total charge on L.H.S. is -2+14(+1) = +12. The total charge on R.H.S. is 2(+3) = +6. Thus, add $6e^-$ on L.H.S.

$$Cr_2O_7^{-2} + 14H^+ \rightarrow 2Cr^{+3} + 7H_2O \text{ ....(1)}$$

Oxidation half reaction:

(i) Balance Mass:

As atoms are already balanced.

$$H_3AsO_3 \rightarrow H_3AsO_4$$

Add 1 H_2O on L.H.S. to balance O atoms.

$$H_3AsO_3 + H_2O \rightarrow H_3AsO_4$$

Add $2H^+$ on R.H.S. to balance H.

$$H_3AsO_3 + H_2O \rightarrow H_3AsO_4 + 2H^+$$

(ii) Balance Charge:

The L.H.S. has no charge. The total charge on R.H.S. is +2. Thus, add $2e^-$ on R.H.S.

$$H_3AsO_3 + H_2O \rightarrow H_3AsO_4 + 2H^+ + 2e^- \text{ ....(2)}$$

Step 3: Balance the loss and gain of electrons in both half reactions and add them.

The reduction half-reaction has $6e^-$ while the oxidation half-reaction has $2e^-$ Thus multiply Eq. (2) by 3 and then add to Eq. (1).

$$Cr_2O_7^{-2} + 14H^+ + 6e^- \rightarrow 2Cr^{+3} + 7H_2O$$
$$3H_3AsO_3 + 3H_2O \rightarrow 3H_3AsO_4 + 6H^+ + 6e^-$$

$$Cr_2O_7^{-2} + 14H^+ + 3H_3AsO_3 + 3H_2O + 6e^- \rightarrow 2Cr^{+3} + 7H_2O + 3H_3AsO_4 + 6H^+ + 6e^-$$

Step 4: Strike out duplication of species on both the sides.

Duplications are $6e^-$, $6H^+$ and $3H_2O$. Cancel this out from both the sides.

$$Cr_2O_7^{-2} + 8H^+ + 3H_3AsO_3 \rightarrow 2Cr^{+3} + 4H_2O + 3H_3AsO_4$$

(i) $Cu + NO_3^{-1} \rightarrow Cu^{+2} + NO$

Step1: Split the reaction into two half reactions.

Reduction half reaction:

$NO_3^{-1} \rightarrow NO$

Oxidation half reaction:

$Cu \rightarrow Cu^{+2}$

Step2. Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

N atoms are already balanced.

$NO_3^{-1} \rightarrow NO$

Add $2H_2O$ on R.H.S. to balance O atoms.

$NO_3^{-1} \rightarrow NO + 2H_2O$

Add $6H^+$ on L.H.S. to balance H atoms.

$NO_3^{-1} + 4H^+ \rightarrow NO + 2H_2O$

(ii) Balance Charge:

The total charge on L.H.S. is -1+4(+1) = +3. The R.H.S. has no charge. Thus, add $5e^-$ on L.H.S.

$NO_3^{-1} + 4H^+ + 3e^- \rightarrow NO + 2H_2O$ ....(1)

Oxidation half reaction

(i) Balance Mass:

Cu atoms are already balanced.

$Cu \rightarrow Cu^{+2}$

(ii) Balance Charge:

The L.H.S. has no charge. The total charge on R.H.S. is +2. Thus, add $2e^-$ on R.H.S.

$Cu \rightarrow Cu^{+2} + 2e^-$ .... (2)

Step3: Balance the loss and gain of electrons in both half reactions and add them.

The reduction half-reaction has $3e^-$ while the oxidation half-reaction has $2e^-$ Thus multiply Eq. (1) by 2, Eq. (2) by 3 and then add the resulting equations.

$2NO_3^{-1} + 8H^+ + 6e^- \rightarrow 2NO + 4H_2O$

$3Cu \rightarrow Cu^{+2} + 6e^-$

$3Cu + 2NO_3^{-1} + 8H^+ + 6e^- \rightarrow 2NO + 4H_2O + 3Cu^{+2} + 6e^-$

Step 4: Strike out duplication of species on both the sides.

Duplications are $6e^-$ Cancel this out from both the sides.

$2NO_3^{-1} + 8H^+ + 3Cu \rightarrow 2NO + 4H_2O + Cu^{+2}$

15. Balance the following equations by oxidation number method

(i) $MnO_2 + HCl \longrightarrow MnCl_2 + H_2O + Cl_2$

(ii) $HNO_3 + HI \longrightarrow NO + I_2 + H_2O$

(iii) $Ag + H_2S + O_2 \longrightarrow Ag_2S + H_2O$

(iv) $Zn + HNO_3 \longrightarrow Zn(NO_3)_2 + NO + H_2O$

(v) $Cu + H_2SO_4 \longrightarrow CuSO_4 + SO_2 + H_2O$

(vi) $Br_2 + NaOH \longrightarrow NaBr + NaBrO_3 + H_2O$

(vii) $HI + H_2SO_4 \longrightarrow I_2 + SO_2 + H_2O$

(viii) $NaClO_2 + Cl_2 \longrightarrow NaCl + ClO_2$

(ix) $Na + H_2O \longrightarrow NaOH + H_2$

Solution:

(i) $MnO_2 + HCl \longrightarrow MnCl_2 + H_2O + Cl_2$

Step1: Assign oxidation number to all the atoms involved in the equation.

$$\overset{+4\ +2(-2)}{MnO_2} + \overset{+1\ -1}{HCl} \rightarrow \overset{+2\ \ 2(-1)}{MnCl_2} + \overset{+2\ -2}{H_2O} + \overset{0}{Cl_2}$$

Step2: Identify the elements undergoing a change in oxidation number

Some Cl atoms change their oxidation state from to -1 in HCl to zero in Cl_2. This is due to loss of + $1e^-$ per Cl atom. This is oxidation. However, some Cl atoms retain their oxidation state at -1 in HCl and $MnCl_2$. So, write HCl twice on L.H.S.

$$HCl + MnO_2 + HCl \rightarrow MnCl_2 + H_2O + Cl_2$$

The Mn atom changes its oxidation state from +4 in MnO_2 to +2 in $MnCl_2$. This is due to gain of $2e^-$ per Mn atom. This is reduction.

Step3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.

$$\overset{(+4)}{MnO_2} + \overset{(-1)}{HCl} \xrightarrow{-1e^-} \overset{+2}{MnCl_2} + \overset{(O)}{Cl_2}$$

$+2e^-$

Step4: Balance the loss and gain of electrons. Use multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $1e^-$ loss by **2** and $2e^-$ gain by **1**.

Multiply the coefficient of HCl by 2 in order to equalize the number of Cl atoms for electrons loss.

$$HCl + MnO_2 + 2HCl \rightarrow MnCl_2 + H_2O + Cl_2$$

These coefficients should not be changed after this.

Step5: Balance the rest of the equation by inspection method.

To balance Cl atoms multiply coefficient of other HCl by 2.

$$2HCl + MnO_2 + 2HCl \rightarrow MnCl_2 + 2H_2O + Cl_2$$

or

$$MnO_2 + 4HCl \rightarrow MnCl_2 + 2H_2O + Cl_2$$

This is the balanced equation.

(ii) $HNO_3 + HI \rightarrow NO + I_2 + H_2O$

Step 1: Assign oxidation number to all the atoms involved in the equation.

$$\overset{+1\ +5\ 3(-2)}{HNO_3} + \overset{+1\ -1}{HI} \rightarrow \overset{+2\ -2}{NO} + \overset{0}{I_2} + \overset{2(+1)\ -2}{H_2O}$$

Step2: Identify the elements undergoing a change in oxidation number

The I atom changes its oxidation state from -1 in HI to zero in I_2. This is due to loss of $1e^-$ per I atom. This is oxidation.

The N atom changes its oxidation state from +5 in HNO_3 to +2 in NO. This is due to gain of $3e^-$ per N atom. This is reduction.

Step3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.

$+3e^-$ (gain)

$$\overset{+5}{HNO_3} + \overset{-1}{HI} \longrightarrow \overset{+2}{NO} + \overset{(O)}{I_2}$$

$+1e^-$ (Lose)

$+3e^-$

$$HNO_3 + \overset{-1}{HI} \longrightarrow NO + I_2$$

$3 \times -1e = -3e^-$

Step4: Balance the loss and gain of electrons by multiplying with small suitable whole numbers. Use these multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $3e^-$ gain by **2** and $2e^-$ loss by **3**.

Multiply the coefficients of HI by 3 and that of I_2 by $\frac{3}{2}$, or better to multiply HI by 6 and I_2 by 3 in order to equalize the number of I atoms for electron loss.

$$HNO_3 + 6HI \rightarrow NO + 3I_2 + H_2O$$

These coefficients should not be changed after this.

Step5: Balance the rest of the equation by inspection method.

To balance oxygen atoms multiply coefficient of HNO_3 by 2, NO by 2 and H_2O by 4.

$$2HNO_3 + 6HI \rightarrow 2NO + 3I_2 + 4H_2O$$

This is the balanced equation.

(iii) $Ag + H_2S + O_2 \rightarrow Ag_2S + H_2O$

Step1: Assign oxidation number to all the atoms involved in the equation.

$$\overset{0}{Ag} + \overset{2(+1)\ -2}{H_2S} + \overset{0}{O_2} \rightarrow \overset{2(+1)\ -2}{Ag_2S} + \overset{2(+1)\ -2}{H_2O}$$

Step2: Identify the elements undergoing a change in oxidation number

The Ag atom changes its oxidation state from zero in Ag to +1 in Ag_2S. This is due to loss of $1e^-$ per Ag atom. This is oxidation.

The O atom changes its oxidation state from zero in O_2 to -2 in H_2O. This is due to gain of $2e^-$ per O atom. This is reduction.

Step3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.

Loss of $-1e^-$

$$\overset{0}{Ag} + \overset{0}{O_2} \longrightarrow \overset{+1}{Ag_2S} + \overset{-2}{H_2O}$$

$+2e^-$ (gain)

Step4: Balance the loss and gain of electrons by multiplying with small suitable whole numbers. Use these multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply 4e⁻ gain by 1 and 1e⁻ loss by 4

$$-1e^- \times 2 = -2e^-$$

$$Ag + O_2 \longrightarrow Ag_2S + H_2O$$

$$+2e^-$$

Multiply the coefficients of Ag by 4 and that of Ag_2S by 2 in order to equalize the number of Ag atoms for electrons loss.

$$4Ag + H_2S + O_2 \rightarrow 2Ag_2S + H_2O$$

Step5: Balance the rest of the equation by inspection method.

To balance S atoms, multiply coefficient of H_2S by 2.

$$4Ag + 2H_2S + O_2 \rightarrow 2Ag_2S + 2H_2O$$

This is the balanced equation.

(iv) $Zn + HNO_3 \rightarrow Zn(HNO_3)_2 + NO + H_2O$

Step1: Assign oxidation number to all the atoms involved in the equation.

$$\overset{0}{Zn} + \overset{+1}{H}\overset{+5}{N}\overset{3(-2)}{O_3} \rightarrow \overset{+2}{Zn}(\overset{+5}{N}\overset{2(-1)}{O_3})_2 + \overset{+2}{N}\overset{-2}{O} + \overset{2(+1)}{H_2}\overset{-2}{O}$$

Step 2: Identify the elements undergoing a change in oxidation number.

The Zn atom changes its oxidation state from zero in Zn to +2 in $Zn(NO_3)_2$. This is due to loss of $2e^-$ per Zn atom. This is oxidation.

Some N atoms change their oxidation state from +5 in HNO_3 to +2 in NO. This is due to gain of $3e^-$ per N atom. This is reduction. Some N atoms retain their oxidation state at +5 in HNO_3 and $Zn(NO_3)_2$. Thus write HNO_3 twice on L.H.S.

$$HNO_3 + Zn + HNO_3 \rightarrow Zn(NO_3)_2 + NO + H_2O$$

Step3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.

$$+3e^-\text{(gain)}$$

$$\overset{0}{Zn} + \overset{+5}{HNO_3} \longrightarrow \overset{+2}{Zn(NO_3)_2} + \overset{+2}{NO}$$

Lose $+2e^-$

Step 4: Balance the loss and gain of electrons by multiplying with small suitable whole numbers. Use these multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $3e^-$ gain by **2** and $2e^-$ loss by **3**.

$$+3e^- \times 2 = 6e^-$$

$$\overset{0}{Zn} + H\overset{+5}{N}O_3 \longrightarrow \overset{+2}{Zn}(NO_3)_2 + \overset{+2}{N}O$$

$$+2e^- \text{ (gain)}$$

Multiply the coefficients of Zn and that of $Zn(NO_3)_2$ by 3.

Multiply coefficients of HNO_3 and NO by 2.

$$HNO_3 + 3Zn + 2HNO_3 \rightarrow 3Zn(NO_3)_2 + 2NO + H_2O$$

Step5: Balance the rest of the equation by inspection method.

Balance equation in the following step-wise manner.

To balance nitrogen atoms multiply coefficient of other HNO_3 by 6.

To balance oxygen atoms multiply coefficient of H_2O by 4.

$$6HNO_3 + 3Zn + 2HNO_3 \rightarrow 3Zn(NO_3)_2 + 2NO + 4H_2O$$

or $$3Zn + 8HNO_3 \rightarrow 3Zn(NO_3)_2 + 2NO + 4H_2O$$

This is the balanced equation.

(v) $Cu + H_2SO_4 \rightarrow CuSO_4 + SO_2 + H_2O$

Step1: Assign oxidation number to all the atoms involved in the equation.

$$\overset{0}{Cu} + \overset{2(+1)}{H_2}\overset{+6}{S}\overset{4(-2)}{O_4} \rightarrow \overset{+2}{Cu}\overset{+6}{S}\overset{4(-2)}{O_4} + \overset{+4}{S}\overset{2(-2)}{O_2} + \overset{2(+1)}{H_2}\overset{-2}{O}$$

Step 2: Identify the elements undergoing a change in oxidation number.

The Cu atom changes its oxidation state from zero in Cu to +2 in $CuSO_4$. This is due to loss of $2e^-$ per Cu atom. This is oxidation.

Some S atom changes its oxidation state from +6 in H_2SO_4 to +4 in SO_2. This is due to gain of $2e^-$ per S atom. This is reduction. Some S atoms retain their oxidation state at +6 in H_2SO_4 and $CuSO_4$. Thus write H_2SO_4 twice on L.H.S.

$$H_2SO_4 + Cu + H_2SO_4 \rightarrow CuSO_4 + SO_2 + H_2O$$

Step3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.

$$\text{gain } +2e^-$$

$$\overset{0}{Cu} + H_2\overset{+6}{S}O_4 \longrightarrow \overset{+2}{Cu}SO_4 + \overset{+4}{S}O_2$$

$$\text{(lost) } -2e^-$$

Step4: Since, loss and gain of electrons is equal, therefore, balance the rest of the equation by inspection method.

To balance O atoms multiply coefficient of H_2O by 2.

$$H_2SO_4 + Cu + H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$$

or $$Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$$

This is the balanced equation.

(vi) $Br_2 + NaOH \rightarrow NaBr + NaBrO_3 + H_2O$

Step1: Assign oxidation number to all the atoms involved in the equation.

$$\overset{0}{Br_2} + \overset{+1}{Na}\overset{-2}{O}\overset{+1}{H} \rightarrow \overset{+1}{Na}\overset{-1}{Br} + \overset{+1}{Na}\overset{+5}{Br}\overset{3(-2)}{O_3} + \overset{2(+1)}{H_2}\overset{-2}{O}$$

Step 2: Identify the elements undergoing a change in oxidation number.

Some Br atoms change their oxidation state from zero in Br_2 to +5 in $NaBrO_3$. This is due to loss of $5e^-$ per Br atom. This is oxidation.

Some Br atoms change their oxidation state from zero in Br_2 to -1 in NaBr. This is due to gain of $1e^-$ per Br atom. This is reduction. So, write Br_2 twice on L.H.S.

$$Br_2 + Br_2 + NaOH \rightarrow NaBr + NaBrO_3 + H_2O$$

Step3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.

(gain) $+1e^-$

$$\overset{0}{Br} + \overset{0}{Br} + \overset{(+1)(-2)(+1)}{NaOH} \longrightarrow \overset{(+1)(-1)}{NaBr} + \overset{(+1)(+5)(-6)}{NaBrO_3} + \overset{(2(+1)-2)}{H_2O}$$

Lost $-5e^-$

Step4: Balance the loss and gain of electrons by multiplying with small suitable whole numbers. Use these multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $2e^-$ gain by **5** and $10e^-$ loss by **1**.

$+ 1e^- \times 5 = +5e^-$

$$\overset{0}{Br} + \overset{0}{Br} + NaOH \longrightarrow NaBr + Na\overset{+5}{Br}O_3 + H_2O$$

$-5e^-$

Multiply the coefficients of Br_2 with 5 and that of NaBr with 10 in order to equalize the number Br atoms for electron gain.

$$Br_2 + 5Br_2 + NaOH \rightarrow 10NaBr + NaBrO_3 + H_2O$$

These coefficients should not be changed after this.

Step5: Balance the rest of the equation by inspection method.

Balance equation in the following step-wise manner

To balance Br atoms multiply coefficient of $NaBrO_3$ by 2.

To balance sodium atoms multiply coefficient of NaOH by 12.

To balance oxygen atoms multiply coefficient of H_2O by 6.

$$Br_2 + 5Br_2 + 12NaOH \rightarrow 10NaBr + 2NaBrO_3 + 6H_2O$$

or $$6Br_2 + 12NaOH \rightarrow 10NaBr + 2NaBrO_3 + 6H_2O$$

or divide by 2, $$3Br_2 + 6NaOH \rightarrow 5NaBr + NaBrO_3 + 3H_2O$$

This is the balanced equation.

(vii) $HI + H_2SO_4 \rightarrow I_2 + SO_2 + H_2O$

Step1: Assign oxidation number to all the atoms involved in the equation.

$$\overset{+1\ -1}{HI} + \overset{2(+1)+6\ 4(-2)}{H_2SO_4} \rightarrow \overset{0}{I_2} + \overset{+4\ 2(-2)}{SO_2} + \overset{2(+1)-2}{H_2O}$$

Step 2: Identify the elements undergoing a change in oxidation number.

The I atom changes its oxidation state from -1 in HI to zero in I_2. This is due to loss of $1e^-$ per I atom. This is oxidation.

The S atom changes its oxidation state from +6 in H_2SO_4 to +4 in SO_2. This is due to gain of $2e^-$ per S atom. This is reduction.

Step3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.

(lost) $-1e^-$

$$\overset{-1}{HI} + H_2SO_4 \longrightarrow \overset{0}{I_2} + \overset{4}{SO_2}$$

(gain) $+2e$

Step4: Balance the loss and gain of electrons by multiplying with small suitable whole numbers. Use these multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $3e^-$ gain by **1** and $1e^-$ loss by **2**

(lost) $-1e^- \times 2 = -2e^-$

$$\overset{-1}{HI} + H_2SO_4 \longrightarrow \overset{0}{I_2} + \overset{4}{SO_2}$$

(gain) $+2e^-$

Multiply the coefficient of HI by 2 so that number of *I* atoms remains equal for electron loss.

$$2HI + H_2SO_4 \rightarrow I_2 + SO_2 + H_2O$$

These coefficients should not be changed after this

Step 5: Balance the rest of the equation by inspection method.

To balance oxygen atoms multiply coefficient of H_2O by 2.

$$2HI + H_2SO_4 \rightarrow I_2 + SO_2 + 2H_2O$$

This is the balanced equation.

(viii) $NaClO_2 + Cl_2 \rightarrow NaCl + ClO_2$

Step1: Assign oxidation number to all the atoms involved in the equation.

$$\overset{+1\ +3\ 2(-2)}{NaClO_2} + \overset{0}{Cl_2} \rightarrow \overset{+1\ -1}{NaCl} + \overset{+4\ 2(-2)}{ClO_2}$$

Step2: Identify the elements undergoing a change in oxidation number.

The Cl atom changes its oxidation state from +3 in $NaClO_2$ to +4 in ClO_2. This is due to loss of $1e^-$ per Cl atom. This is oxidation.

The Cl atom changes its oxidation state from zero in Cl_2 to -1 in NaCl. This is due to gain of $1e^-$ per Cl atom. This is reduction.

Step3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.

(lose) $-1e$

$$Na\overset{+3}{Cl}O_2 + \overset{(0)}{Cl_2} \longrightarrow Na\overset{(-1)}{Cl} + \overset{+4}{Cl}O_2$$

Gain $+1e^-$

Step 4: Balance the loss and gain of electrons by multiplying with small suitable whole numbers. Use these multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $2e^-$ gain by **1** and $1e^-$ loss by 2.

Multiply the coefficients of $NaClO_2$ and that of ClO_2 by 2, so that number of Cl atoms remains equal for electron loss

$$2NaClO_2 + Cl_2 \rightarrow NaCl + 2ClO_2$$

These coefficients should not be changed after this

Step 5: Balance the rest of the equation by inspection method.

To balance sodium atoms multiply coefficient of NaCl by 2

$$2NaClO_2 + Cl_2 \rightarrow 2NaCl + 2ClO_2$$

This is the balanced equation.

(ix) $Na + H_2O \rightarrow NaOH + H_2$

Step 1: Assign oxidation number to all the atoms involved in the equation.

$$\overset{0}{Na} + \overset{2(+1)-2}{H_2O} \rightarrow \overset{+1\ -2\ +1}{NaOH} + \overset{0}{H_2}$$

Step 2: Identify the elements undergoing a change in oxidation number.

The Na atom changes its oxidation state from zero in Na to +1 in NaOH. This is due to loss of $1e^-$ per Na atom. This is oxidation.

The H atom changes its oxidation state from +2 in H_2O to zero in H_2. This is due to gain of $1e^-$ per H atom. This is reduction

Step3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.

(gain) $+1e^-$

$$\overset{(0)}{Na} + \overset{2(+1)(-2)}{H_2O} \longrightarrow \overset{(+1)(-2)(+1)}{NaOH} + \overset{(0)}{H_2}$$

(lose) $-1e$

Step 4: Balance the loss and gain of electrons by multiplying with small suitable whole numbers. Use these multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $2e^-$ gain by 1 and $1e^-$ loss by 2.

Multiply the coefficients of Na and that of NaOH by 2, so that number of Na atoms equalize for electron loss.

Multiply coefficients of HNO_3 and NO by 2.

$$2Na + H_2O \rightarrow 2NaOH + H_2$$

These coefficients should not be changed after this

Step 5: Balance the rest of the equation by Inspection method.

To balance oxygen atoms multiply coefficient of H_2O by 2.

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

This is the balanced equation

16. A dead lead storage battery has 5g of $PbSO_4$ deposited on each of its anodes.It is connected to a charger that supplies 0.12 ampere of current at a voltage of 13V. To which electrode, Pb or PbO_2 should the anode of charger be connected? Write the half reaction occuring at this electrode during charging?

Ans: Recharging a dead lead storage battery:

i. Battery can be recharged by connecting its anode to the negative terminal of direct current and the cathode to the positive terminal of the direct current

ii. Reverse chemical reactions occur at anode and cathode of the battery.

iii. Thus deposition of Pb on anode and PbO_2 on cathode takes place. In this way we can recharge the battery.

Redox reactions during recharging:

The reactions of recharging of battery are as follows.

At cathode:

$$PbSO_{4(S)} + 2\bar{e} \longrightarrow Pb_{(S)} + SO_4^{-2}{}_{(aq)}$$

At anode:

$$PbSO_{4(S)} + 2H_2O_{(l)} \longrightarrow PbO_{2(S)} + SO_4^{-2}{}_{(aq)} + 4H^+_{(aq)} + 2\bar{e}$$

Net reaction: $2PbSO_{4(S)} + 2H_2O_{(l)} \longrightarrow Pb_{(S)} + PbO_{2(S)} + 4H^+_{(aq)} + 2SO_4^{-2}{}_{(aq)}$

After recharging H_2SO_4 solution is concentrated again bringing density to its initial value of 1.25 g cm^{-3}

17. Calculate the mass of Cu and O_2 produced by the electrolysis of $CuSO_4$ solution on passing 5.0 amperes of current for 2 hrs. What would be the volume of O_2 at S.T.P. Following reactions occurs at the electrodes.

At Anode $2H_2O_{(l)} \longrightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^-$

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

(Ans: 0.1865g, 2.089dm^3)

Solution:

Calculation of total charge passed:

Time = 2 hours = 2 × 60 × 60 = 7200 s

Current = 5 A

As we know that

$$\text{Current} = \frac{\text{Charge}}{\text{Time (s)}}$$

So Charge = Current × Time (s)

= 5 A × 7200 s

= 36000 C

As

96500 C = 1F

$$36000\ C = \frac{1F}{96500\ C} \times 36000\ C = 0.3731\ F$$

In the electrolysis of molten bauxite, the cathode reaction is

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

It shows that for every 2 moles of electrons, 1 mole of Cu is produced.

1 F = 1 mole of electron.

2 F = 2 moles of electrons = 1 mole of Cu.

Thus,

2 F charge produces Cu = 1 mole

0.3731 F charge produces Cu = $\frac{1}{2} \times 0.3731$

= 0.1866 moles

Atomic mass of Cu = 63.5 g $mole^{-1}$

So

1 mole of Cu = 63.5 g

0.1866 moles of Cu = 63.5×0.1866

= 11.849 g

The reaction at anode is oxidation of O^{2-} ions.

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

It shows that for every 4 moles of electrons, 1 mole of O_2 is produced.

1F = 1 mole of electron

4 F = 4 moles of electrons = 1 mole of O_2.

Thus ,

4 F electricity produce O_2 = 1 mole

0.3731 F of electricity produce O_2 = $\frac{1}{4} \times 0.3731$

= 0.0933 moles of O_2

Molecular mass of O_2 = 32 g $mole^{-1}$

Hence

1 mole of O_2 = 32 g

0.0933 moles of Cu = 32×0.0933

= 2.986 g

Thus , at S.T.P.

1 mole of O_2 = 22.414 dm^3

0.0933 moles of O_2 = $0.0933 \times 22.414\ dm^3$

= 2.091 dm^3

18. If a chromium steel bicycle handlebar is scratched, exposing steel, will chromium or steel corrode?

Ans: Chromium is ranked higher in activity series than iron (steelis a alloy of iron) so it has more activity than iron and it will react more radially with atospheric oxygen hence the chromium will corrode. Thus if a chromium steel bicycle handlebar is scratched, exposing steel, will chromium will corrode instead of steel .

19. Use the activity series of metals to predict the products of following single replacement reactions. Give reason.

i) $Cu(s) + HCl(aq) \longrightarrow$

ii) $Fe_2O_3(s) + H_2(g) \longrightarrow$

iii) $Mg(s) + CrCl_3(s) \longrightarrow$

iv) $NiO(s) + H_2(g) \longrightarrow$

v) $Cd(s) + CuSO_4(aq) \longrightarrow$

Ans: In active series a metal which is ranked higher can displace low. ranked metal from its aqueous solution.

(i) $Cu_{(s)} + HCl_{(aq)} \longrightarrow$

Cu is ranked lower in activity series than hydrogen. Thus, it will not displace hydrogen from HCl.

$Cu_{(s)} + HCl_{(aq)} \longrightarrow$ No Reaction

(ii) $Fe_2O_{3(s)} + H_{2(g)} \longrightarrow$

The hydrogen is ranked below *Fe* in activity series, hence it cannot displace *Fe*.

$$2Fe_2O_{3(s)} + 3H_{2(g)} \longrightarrow \text{No reaction}$$

(iii) $Mg_{(s)} + CrCl_{3(s)} \rightarrow$

Mg is ranked higher than *Cr* in activity series thus it can displace *Cr*. Thus this reaction will occur

$$3Mg_{(s)} + 2CrCl_{3(s)} \longrightarrow 3MgCl_2 + 2Cr$$

(iv) $NiO_{(s)} + H_{2(g)} \longrightarrow$

The hydrogen is ranked below *Ni* in activity series thus it cannot displace *Ni*.

$NiO_{(s)} + H_{2(g)} \longrightarrow$ No Reaction

(v) $Cd_{(s)} + CuSO_{4(aq)} \longrightarrow$

The Cd is ranked above Cu in activity series, hence it will displace Cu. Thus, this reaction will occur.

$$Cd_{(s)} + CuSO_{4(aq)} \longrightarrow CdSO_4 + Cu$$

20. Justify that photosynthesis in plants is a redox reaction.

$$6CO_2 + 6H_2O \xrightarrow{chlorophyll} C_6H_{12}O_6 + 6CO_2$$

Solution:

Assigning oxidation numbers:

$$6C\overset{+4}{}O_2^{\,2(-2)} + 6H_2^{\,2(+1)}O^{-2} \xrightarrow{chlorophyll} C_6^{\,6(0)} H_{12}^{\,12(+1)} O_6^{\,6(-2)} + 6O_2^{\,0}$$

(i) Oxidation.

The *O* atom changes its oxidation state from -2 to zero in O_2. This is due to loss of electrons. Hence it is oxidation.

(ii) Reduction.

The *C* atom changes its oxidation state from +4 to zero in $C_6H_{12}O_6$. This is due to gain of electrons. Hence this is reduction.

Conclusion:

As oxidation and reduction occur in photosynthesis therefore it is a redox reaction.

21. Metal A can displace metal B from its aqueous solution, metal B can displace metal D from its aqueous solution, but it cannot displace metal C from its aqueous solution. Metal C can displace metal A from its aqueous solution. Metals B and D can be displaced from their aqueous solution by hydrogen. Arrange these metals and hydrogen in order of decreasing activity.

Solution:

Given information:

i. Metal A can displace metal B from its aqueous solution. It means that metal A is more active than metal B

Metal A > Metal B

ii. Metal B can displace metal D from its aqueous solution. It means that metal B is more active than metal D

Metal A > Metal B > Metal D

iii. Metal C can displace metal A from its aqueous solution. It means that metal C is more active than metal A.

Metal C > Metal A > Metal B > Metal D

iv. Metal B and D can be displaced by hydrogen from their aqueous solution. Thus, it means that metal B and D are less active than hydrogen.

Conclusion:

Arrange these metals and hydrogen in order of decreasing activity

Metal C > Metal A > Hydrogen > Metal B > Metal D

22. Differences in the standard potentials of electrodes A and B, B and C are 1.5V and 3.0V respectively.

a) **Calculate the cell potential of the working cell between A and C.**

b) **Identify cathode in this cell.**

Solution:

(a) Calculate the cell potential of the working cell between A and C.

Given that

$E_B^o - E_A^o = 1.50\ V$ ....(i)

$E_B^o - E_C^o = 3.0\ V$ ....(ii)

For the calculate the cell potential of the working cell between A and C. we have to Subtract equation (i) from equation (ii)

$$(E_B^o - E_C^o) - (E_B^o - E_A^o) = 3.0 - 1.50$$
$$E_B^o - E_C^o - E_B^o + E_A^o = 1.50\ V$$
$$E_A^o - E_C^o = 1.50\ V$$

(b) Identify cathode in this cell.

$$E_A^o = 1.50\ V + E_C^o$$

As the standard electrode potentioal of A is greater than B therefore A will act as cathode and B will act as anode.

23. Standard electorde potentials of three electrodes are in the following order,

$$E_A^0 > E_B^0 < E_C^0$$

Identify anode in the working cell, when (a) Electrodes A and C are joined in series. (b) Electrodes B and C are joined in series (c) Electrodes A and B are joined in series.

Ans: The emf measured when a metal/metal ion electrode is coupled to a hydrogen electrode under standard conditions is known as the standard electrode potential of that metal / metal ion combination.

Condition:

$$\text{As } E_A^0 > E_B^0 < E_C^0$$

Principle:

The electrode which has higher standard electrode potential will act as cathode and the electrode which has lower standard electrode potential will act as anode.

The above equation indicates that standard electrode potential of electrode B is less than that of both A and C electrodes.

(a) Electrodes A and C are joined in series.

As no information is provided about the electrode potential of electrode A and C relative to one another. Therefore, we cannot identify anode in this case.

(b) Electrodes B and C are joined in series.

When Electrodes B and C are joined in series electrode B will act as anode and electrode C will act as cathode. We conclude this with the help of following relation.

Since, $E_B^0 < E_C^0$

Therefore,

(c) Electrodes A and B are Joined in series.

As we know that

$$E_A^0 > E_B^0$$

Thus, when electrodes A and B are joined in series electrode B will act as anode and electrode A will act as cathode because the electrode which has higher standard electrode potential will act as cathode and the electrode which has lower standard electrode potential will act as anode.

24. Evaluate the importance of standard reduction potentials.

Ans: i. A reduction potential measures the tendency of a molecule to be reduced by taking up new electrons.

ii. The standard reduction potential is the reduction potential of a molecule under specific, standard conditions.

iii. Standard reduction potentials can be useful in determining the directionality of a reaction.