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?

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 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.

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

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

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

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

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 "cactic element undergoing oxidation or reduction undergo a charge in the Oxidation number.

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

Ans: Oxidation Number:

The Oxidation number, oxidation stare 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 H2. S in S8 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 OF2)

Example 12.1:

Calculate the Oxidation number of Mn in $KMnO_4$

Solution: Oxidation number ...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 $K_2Cr_2O_7$

Solution:

K_2 Cr_2 O_7

(+1)2 + (x)2 + (-2) 7

+2 + 2x + 14 =0

+ 2x + -12 = 0

2x =12

x= 12/2 = 6

Thus oxidation number of $K_2Cr_2O_7$ 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

x = +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 increase in oxidation number (a loss of electrons)

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

Explanation:

Consider the following reaction

…2NaCl(s)

…atoms involved in this reaction and write it…

$$\overset{2}{2Na(s)} + \overset{0}{cl_2(g)} \rightarrow \overset{+1-1}{2NaCl}\ (s)$$

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

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

On the other hand each Cl atom in Cl_2 molecule changes is 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.

$$\overset{0}{Cl_2(g)} + 2e \rightarrow 2Cl^{-1}$$

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.

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

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

4+4 0 ... .

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

The C changes as oxidation number from -4 in CH_4 to +4 in $CO_2(g)$,This is 8 electrons loss. This means C undergoes an increase in oxidation number. On the other hand, C changes its oxidation number from zero in O_2 to -2 in H_2O, 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.

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

At Cathode, $2NH_4^+ + 2MnO_2 + 2e^- \rightarrow Mn_2O_3 + 2NH_3 + 2H_2O$

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

...

In this reaction ...

This means that each...

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.

$N_2(g)+3H_2(g)\rightarrow 2NH_3(g)$

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

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

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

Assigning oxidation number:

0 0 -3+3(1)

$N_2(g)+3H_2(g)\rightarrow 2NH_3(g)$

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.

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

Assigning oxidation number:

0 0 2(+3) 3(-2)

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

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

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

Assigning oxidation number:

0 12(+1) 6(-2) 0 ...2(-2) 2(-1)-2

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

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+ $HNO_3+H_2O\rightarrow H_3PO_4+NO$

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

0 +1 +5 (-2) 3 (+1)2 -2 31+1) +5 (-2)4 2 2

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

Step 2: 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 charge. 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.

$$+3\bar{e}$$

$$\overset{0}{P} + H\overset{+5}{N}O_3 \longrightarrow H_3\overset{+5}{P}O_4 + \overset{-2}{N}O$$

$$-5\bar{e}$$

Step 4: 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.

$$+3\bar{e} \times 5 = +15\bar{e}$$

$$\overset{0}{P} + H\overset{+5}{N}O_3 \longrightarrow H_3\overset{+5}{P}O_4 + \overset{-2}{N}O$$

$$-5\bar{e} \times 3 = -15\bar{e}$$

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

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

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

Step 5: 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 $+ 5HNO_3 + H_2O \rightarrow 3H_3PO_4 + 5NO$

Inspect the equation, it is balanced.

SELF-CHECK EXERCISE 12.3

Using the oxidation number method balance the following equation.

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

Solution:

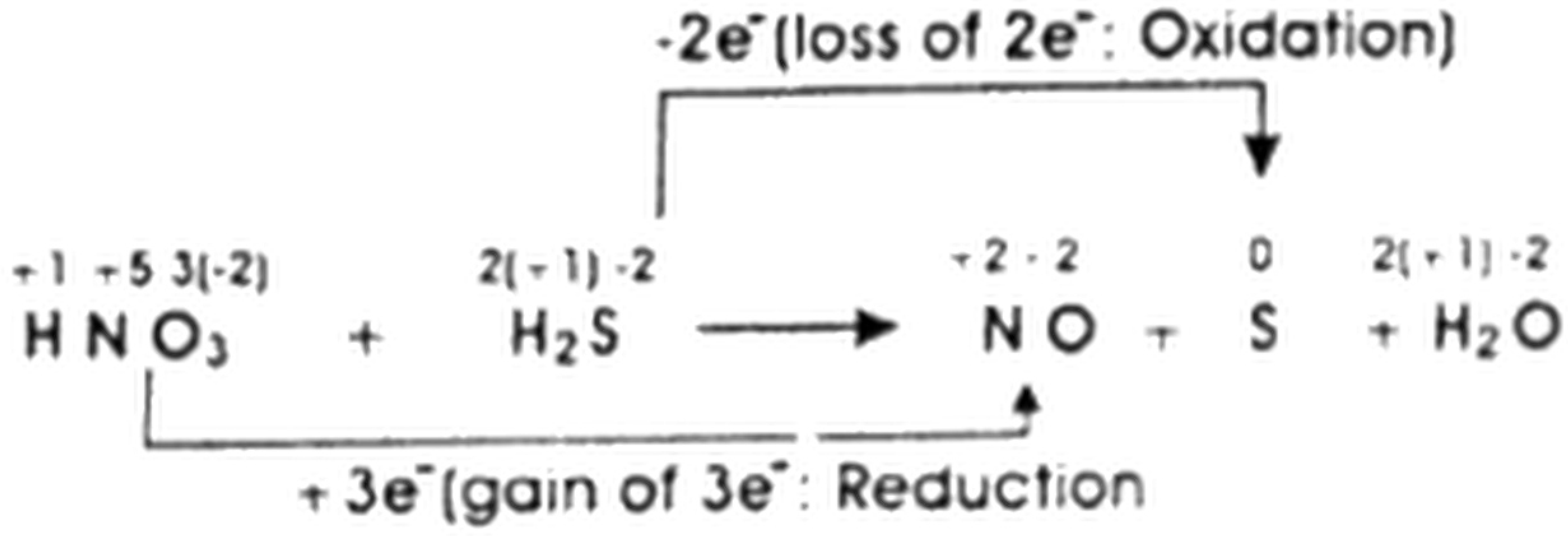
Step 1: Assign oxidation number to all the atoms involved in the equation.

+1 -5 3(-2) 21-1)-2 +2-20 2(-1)-2

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

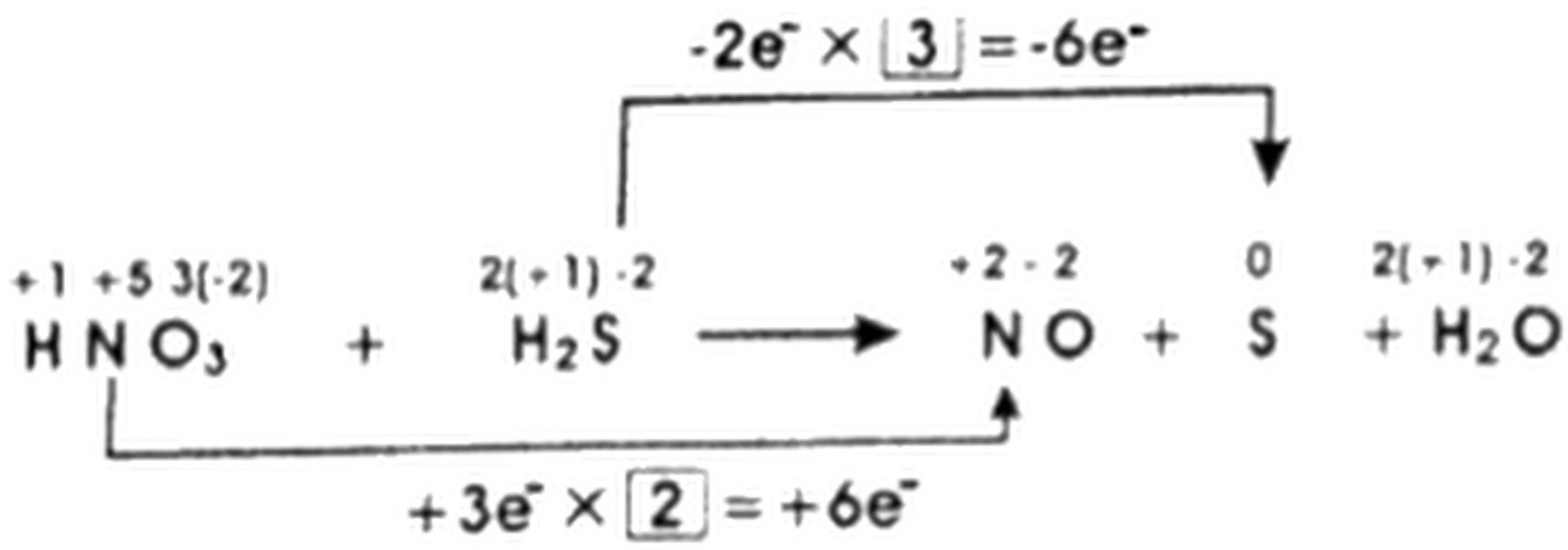
Step 2: Identify the elements undergoing a change in oxidation number N atom changes its oxidation state from +5 to +2 in NO. It is because of 3 electron change The S atom changes its oxidation state from 2 to zero in S It is because of the 2 electron change.

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.



Step 4: 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



Multiply the coefficients of H.S and that of S by 3 Whereas multiply coefficients of HNO_3

and NO by 2

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

These coefficients should not be changed after this

Step 5: 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

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

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.

$[Mn\overset{(-2)4}{O}_4]^{-1} + Fe^{+2} \xrightarrow{acid} Fe^{+3} + Mn^{+2}$

Step 2: 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}

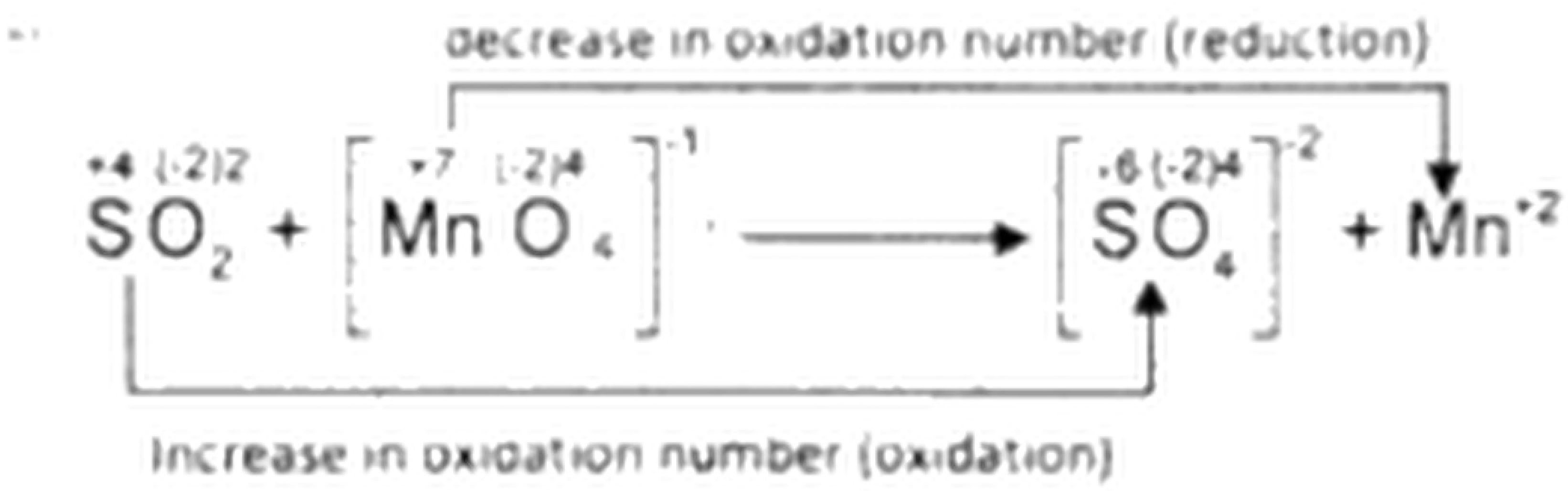
$Fe^{+2} \longrightarrow Fe^{+3}$ (Oxidation half-reaction)

Example 12.7:

Split the following reaction into half-reactions.

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

Solution:



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:

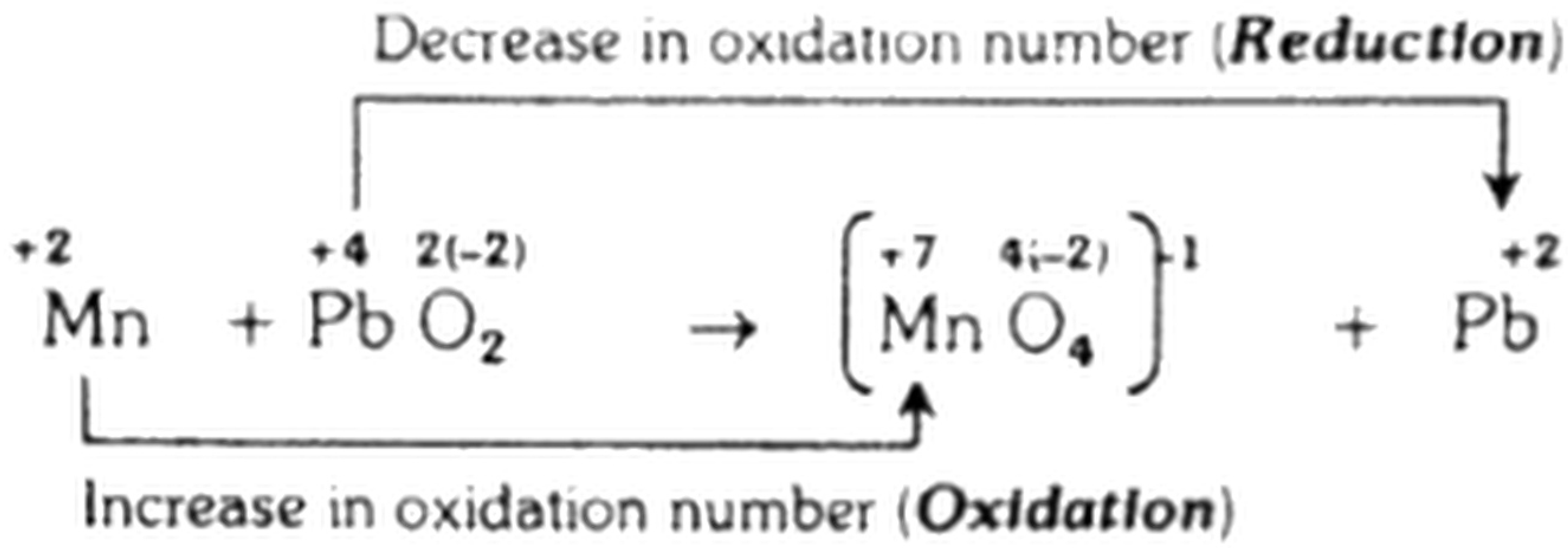
(i) $Mn^{+2} + PbO_2 \rightarrow MnO_4^{-1} + Pb^{+2}$

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

(iii) $NO + MnO_4 \longrightarrow NO_2 + MnO_2$

Solution:

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



Reduction half reaction:

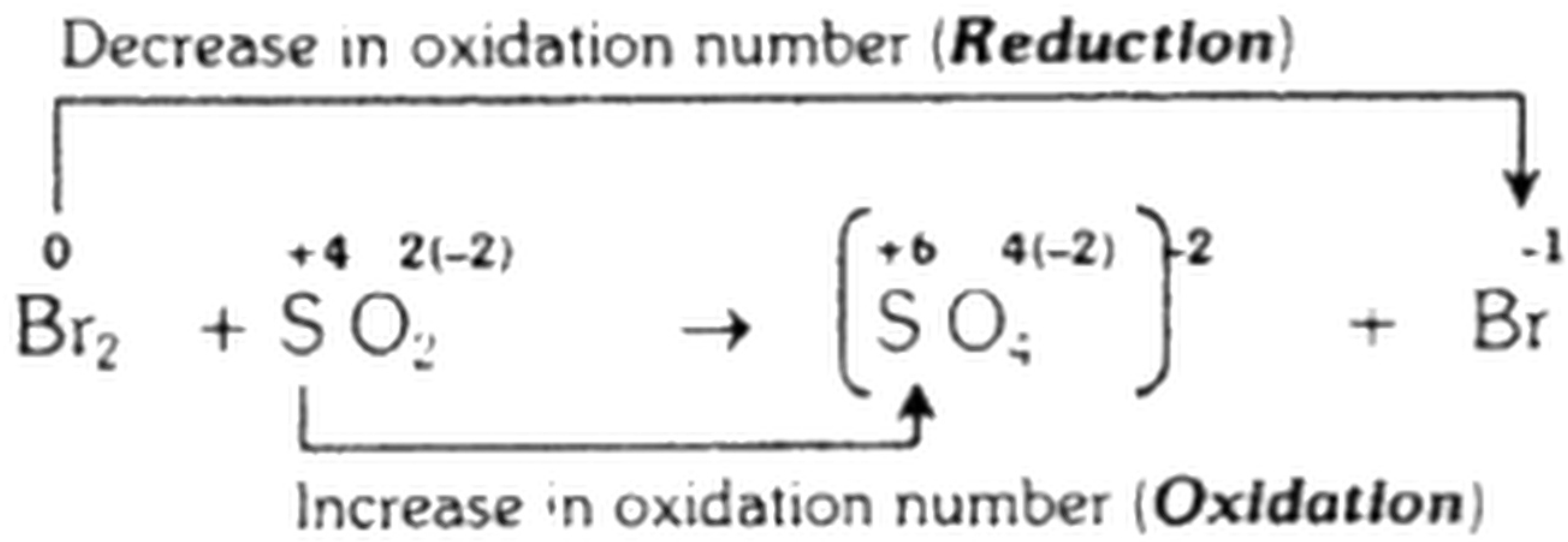
$PbO_2 \rightarrow Pb^{+2}$

Oxidation half reaction:

$Mn^{+2} \rightarrow MnO_4^{-1}$

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

Solution:



Reduction half reaction:

$Br_2 \longrightarrow Br^{-1}$

Oxidation half reaction:

$SO_2 \longrightarrow SO_4^{-2}$

(iii) $NO + MnO_4^{-1} \longrightarrow NO_2^{-1} + MnO_2$

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:

$NO \longrightarrow NO_2^{-1}$

Self-Check Exercise 12.4

Split the following reactions into two half reactions:

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

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

(iii) $N0 + MnO_4^{-1} \rightarrow NO_4^{-1} + MnO_2$

Solution:

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

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:

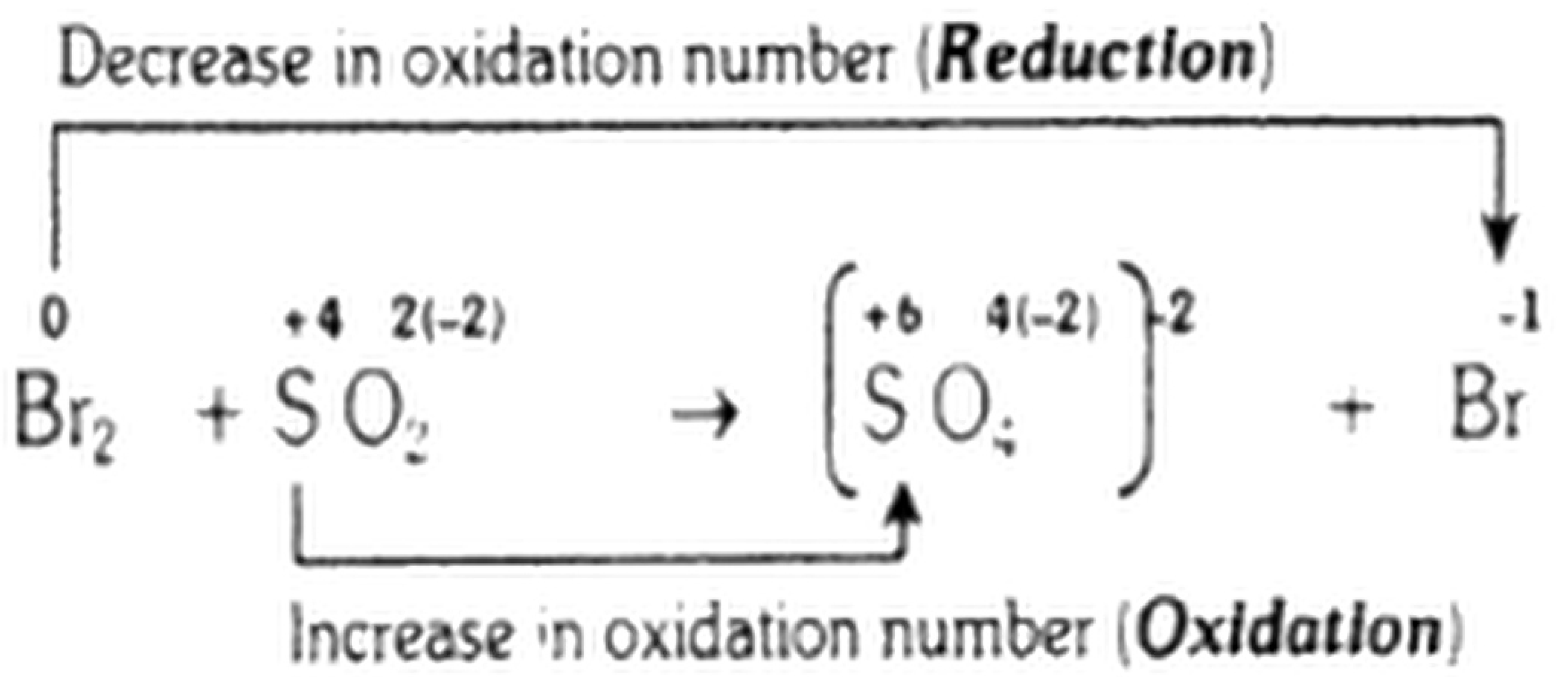
$PbO_2 \rightarrow Pb^{+2}$

Oxidation half reaction:

$Mn^{+2} \rightarrow MnO_4^{-1}$

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

Solution:



Reduction half reaction:

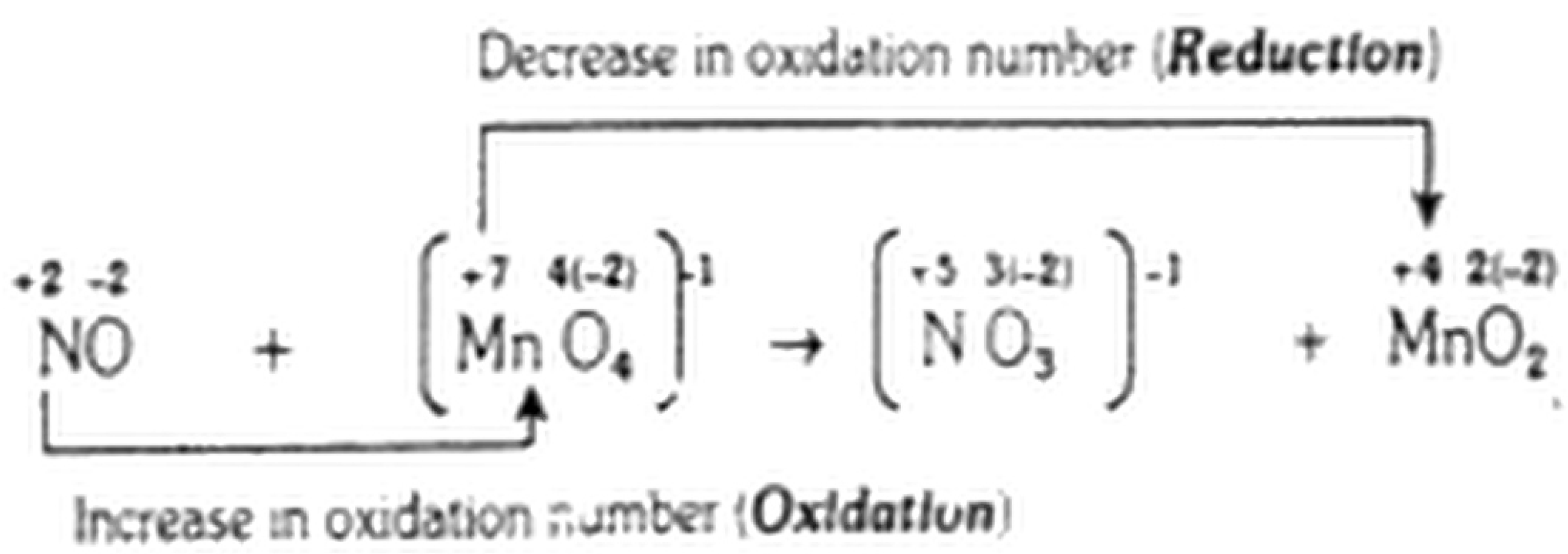
$Br_2 \rightarrow Br^{-1}$

Oxidation half reaction:

$SO_2 \rightarrow SO_4^{-2}$

(iii) $NO + MnO_4^{-1} \quad \rightarrow NO_4^{-1} + MnO_2$

Solution:



Reduction half reaction:

$MnO_4^{-1} \rightarrow MnO_2$

Oxidation half reaction:

$NO \rightarrow NO_3^{-1}$

Q.9How can we balance are dox reaction with the help of half reaction method?

Ans: Half Reaction Method to Balance a Redox Reaction:

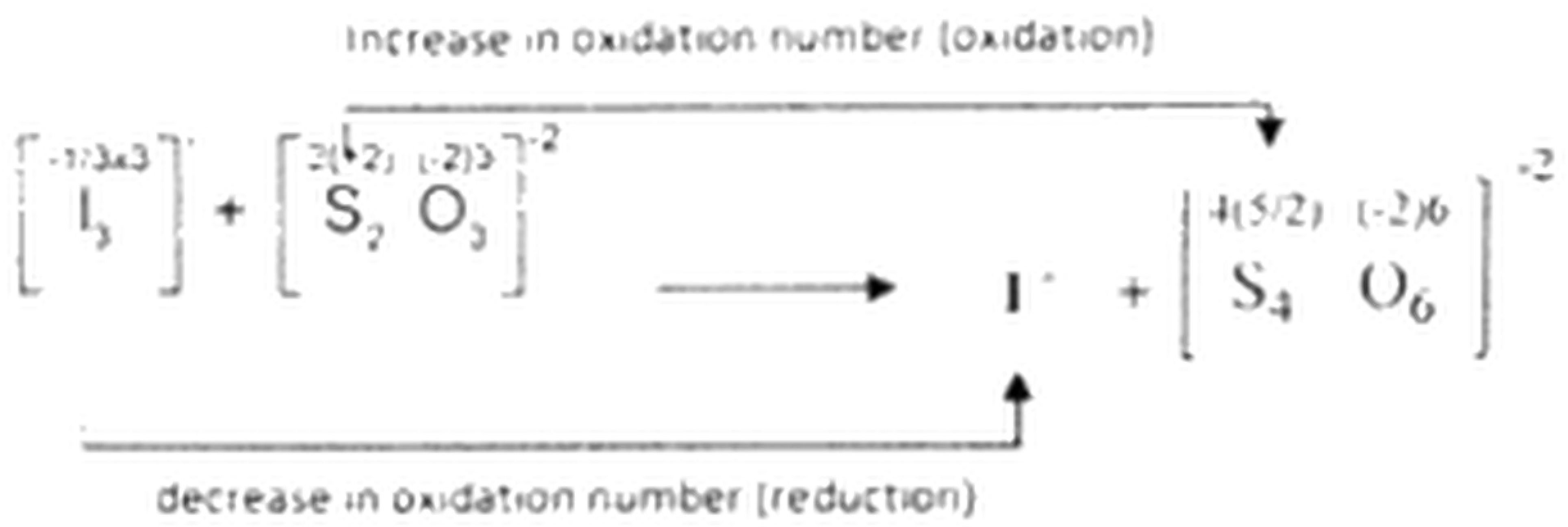
A powerful technique for balancing redox reactions involves dividing these reactions into separate oxidation and reduction half reactions.

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.

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

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



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

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

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.

$I_3^- \rightarrow I^-$

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 sides.

$$I_3^- \rightarrow 3I^-$$

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 $2e^-$

$$I_3^- + 2e^- \rightarrow 3I^- \qquad \ldots\ldots\ldots\ldots\ldots(1)$$

Now balance oxidation half reaction.

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

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

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

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 $2e^-$.

$$2S_2O_3^{-2} \rightarrow S_4O_6^{-2} + 2e^- \qquad \ldots\ldots\ldots\ldots\ldots\ldots\ldots(2)$$

Notice that both the half reactions 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)

$$I_3 + 2e \rightarrow 3I$$

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

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

Step4: Cancel Duplication of Species on both the sides.

Duplications are $2e$ strike these out from both the sides.

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

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.

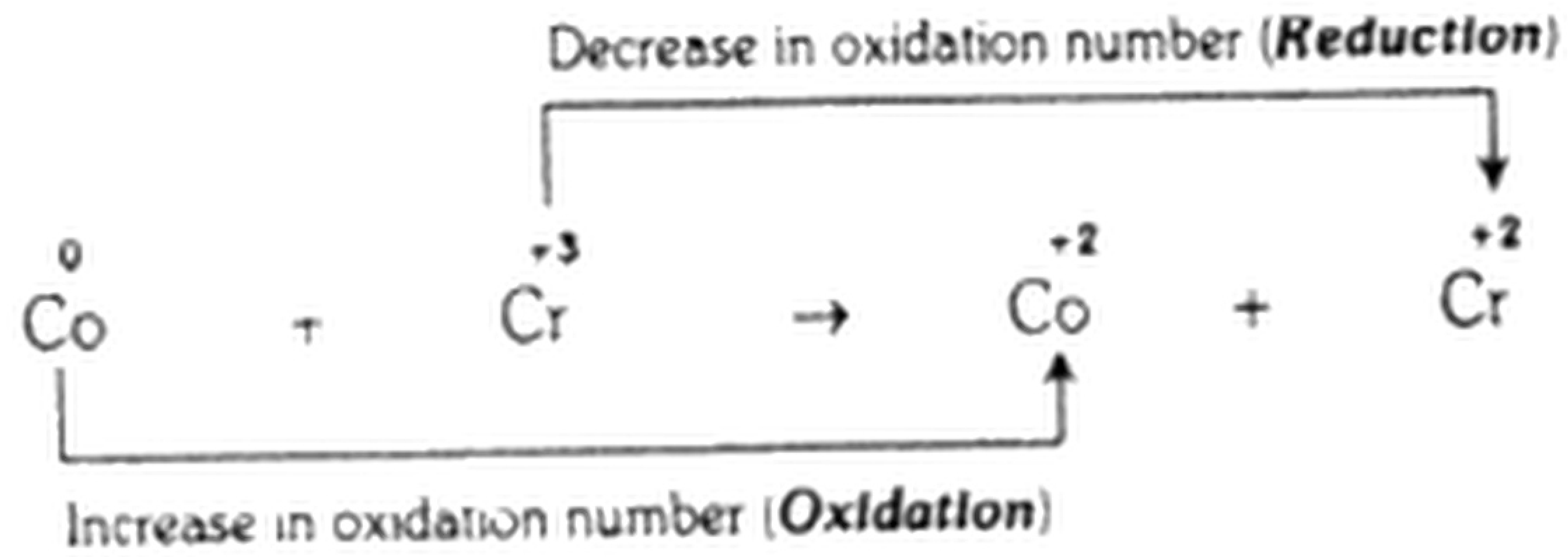
(i) $CO + Cr^{+3} \rightarrow CO^{+2} + Cr^{+2}$

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

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

Solution: **(i)** $CO + Cr^{+3} \rightarrow CO^{+2} + Cr^{+2}$

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



Reduction half-reaction:

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

Oxidation half-reaction:

$Co \rightarrow Co^{+2}$

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

First consider the reduction half-reaction.

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

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 $1e^-$

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

Oxidation half-reaction:

$Co \rightarrow Co^{+2}$

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.

$Co \rightarrow Co^{+2} + 2e^-$ ....(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).

$2Cr^{+3} + 2e^- \rightarrow 2Cr^{+2}$

$Co \rightarrow Co^{+2} + 2e^-$

$Co + 2Cr^{+3} + 2e^- \rightarrow Co^{+2} + 2Cr^{+2} + 2e^-$

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

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

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

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

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

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

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

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

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

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

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

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

Reduction half-reaction:

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

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 $3e^-$.

$$Au^{+3} + 3e^- \rightarrow Au \qquad \text{.........................(i)}$$

Oxidation half reaction:

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

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}.

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

Balance charge:

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

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

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.

$$2Au^{+3} \quad + \quad 6e^- \quad \rightarrow \quad 2Au$$

$$3H_2 \quad \rightarrow \quad 6H^{+1} \quad + \quad 6e^-$$

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

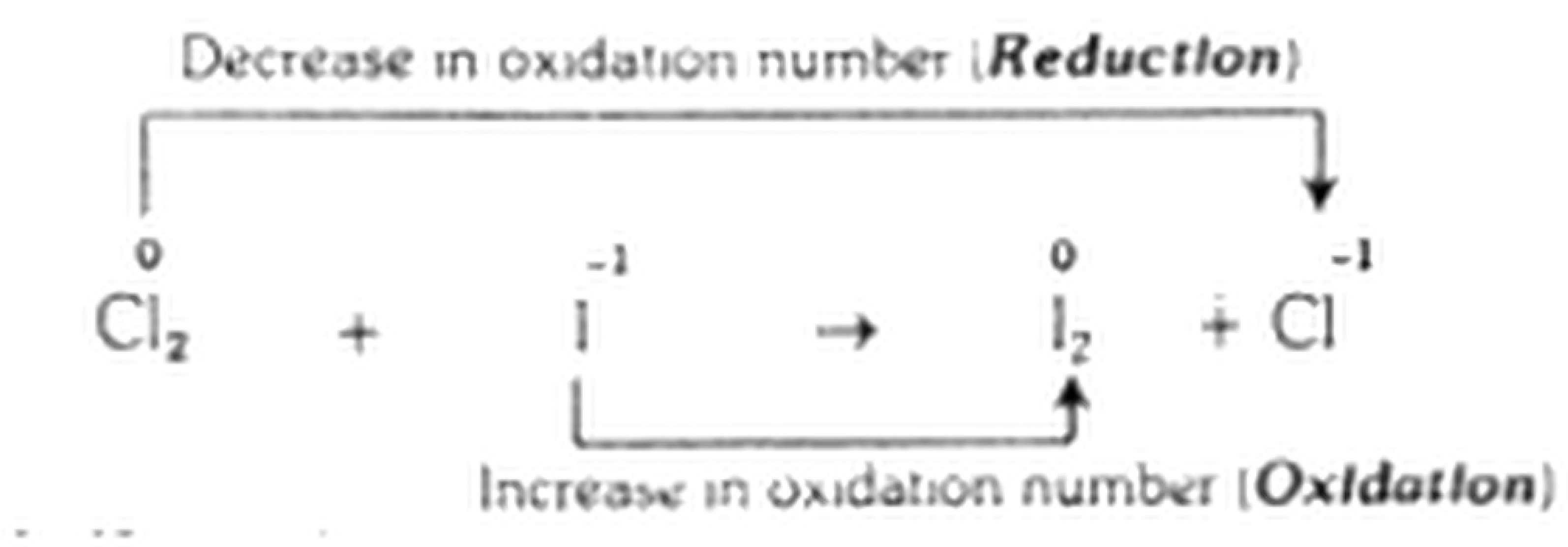
Step 4: Cancel Duplication of Species on both the sides.

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

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

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

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



Reduction half-reaction:

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

Oxidation half-reaction:

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

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:

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

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.

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

Balances Charge:

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

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

Oxidation half-reaction:

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

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

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

Balance Charge:

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

$$2I^{-1} \rightarrow I_2 + 2e^- \quad \text{...(i)}$$

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).

$$Cl_2 \quad + \quad 2e^- \quad \rightarrow \quad 2Cl^{-1}$$

$$2I^{-1} \quad + \quad I_2 + 2e$$

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

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

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

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

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.

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

Solution:

Step1: Split the reaction into two half reactions.

Reduction half reaction:

$$C_{r_2}O_7^{-2} \rightarrow Cr^{+3}$$

Oxidation half reaction:

$$H_2SO_3 \rightarrow HSO_4$$

Step2: Balance each half reaction. First consider reduction half reaction. Two Cr atoms on the left require 2 before Cr^{+3}

$$C_{r_7}O_7^{-2} \rightarrow 2Cr^{+3}$$

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

$$C_{r_7}O_7^{-2} \rightarrow 2Cr^{+3} + 7H_2O$$

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

$$C_{r_7}O_7^{-2} + 14H^+ \rightarrow 2Cr^{+3} + 7H_2O$$

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 $6e^-$.

$$C_{r_7}O_7^{-2} + 14H^+ + 6e^- \rightarrow Cr^{+3} + 7H_2O \qquad \ldots\ldots\ldots\ldots(1)$$

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

$$H_2SO_3 \rightarrow HSO_4^-$$

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.

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

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

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

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

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

Thus we will add $2e$ to the right side

$$H_2SO_3 + H_2O \rightarrow HSO_4^- + 3H^+ + 2e \qquad \ldots\ldots\ldots\ldots\ldots(2)$$

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 produces 2e. Therefore multiplying equation (1) by one and equation (2) by three and adding two equations we get.

$$C_{r_7}O_7^{-2} + 14H^+ + 6e^- \quad \rightarrow \quad 2Cr^{+3} + 7H_2O$$

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

$$C_{r_7}O_7^{-2} + 14H^+ + 3H_2SO_3 + 3H_2O + 6e^- \rightarrow 2Cr^{+3} + 7H_2O + 3HSO_4^- + 9H^+ + 6e^-$$

Step4: Cancel the duplication. Duplications are $\boldsymbol{6e^-}$. $\boldsymbol{3H_2O}$ and $\boldsymbol{9H^+}$. Strike these out from both sides.

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

Example 12.9:

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

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

Solution:

Step1: Split the reaction into two half reactions.

Reduction half reaction:

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

Oxidation half reaction:

$$H_2O_2 \rightarrow O_2$$

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

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

Mn atoms on both the sides 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.

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

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

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

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 $5e^-$.

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

Now consider oxidation half reaction.

$$H_2O_2 \rightarrow O_2$$

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.

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

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

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

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.

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

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

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

Step5: Strike out duplications from both the sides

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

SELF-CHECK EXERCISE 12.6

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

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

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

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

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

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

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

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

d) $Mn0_4^- + C_2O_4^{-2} \rightarrow Mn^{+2} + CO_2$

Solution:

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

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

Step 1: N atoms on the both sides 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.

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

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

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

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 $1e^-$

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

This is the balanced reduction half reaction

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

Reduction half reaction:

Step 1: There is one I atom of the left side and 2 on the right sides so we will balance this by multiplying IO_3^- with 2

$$2IO_3 \rightarrow I_2$$

There are 6 oxygen atoms in the left side and none on the right side so we will add $6H_2O$ on the right side.

$2IO_3 \rightarrow I_2 + 6H_2O$

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

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

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 $10e^-$

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

This is balanced reduction half reaction.

c) $Cl_4^- \rightarrow ClO_3^-$

Step 1: Cl atoms on the both sides are already balanced.

$$Cl_4^- \rightarrow ClO_3^-$$

There are 4 oxygen atoms on the left side and 3 on the right side so we will add 1 H_2O on the right side.

$$Cl_4^- \rightarrow ClO_3^- + H_2O$$

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

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

Step 2:

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

Now balance charges there is one 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 $2e^-$.

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

This is the balanced reduction half reaction.

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

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

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

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

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

Step 2: Balance each half-reaction separately.

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

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

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

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

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 $3e^-$.

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

Oxidation half-reaction:

Balance Mass: $\boldsymbol{Br^- \rightarrow Br_2}$

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

$$2Br^- \rightarrow Br_2$$

Balance Charge:

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

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

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.

$$2NO_3^- + 8H^+ + 6e^- \rightarrow 2NO + 4H_2O$$

$$6Br^- \rightarrow 3Br_2 + 6e^-$$

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

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

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

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

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

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

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

Oxidation half-reaction: $H_3AsO_3 \rightarrow H_3AsO_4$

Step 2: Balance each half-reaction separately.

Reduction half-reaction:

Ce atoms are already balanced.

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

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.

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

Oxidation half-reaction:

Atoms are already balanced.

$$H_3AsO_3 \rightarrow H_3AsO_4$$

To balance oxygen add 1 H_2O on left side.

$H_3AsO_3 \rightarrow H_3AsO_4$

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

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

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 $2e^-$.

$H_3AsO_3 + H_2O \rightarrow H_3AsO_4 + 2H^+ + 2e^-$ ....(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 Eq. (ii).

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

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

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

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

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

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

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

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

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

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

Step 2: Balance each half-reactions separately.

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

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

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 $2e^-$.

$S_2O_8^{-2} + 2e^- \rightarrow 2S_4^{-2}$……………………………………………(i)

Oxidation half-reaction:

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

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

AddH_2O on left side to balance O atoms.

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

Add $14H^+$ on the right side to balance H.

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

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.

$2Cr^{+3} + 7H_2O \rightarrow Cr_2O_7^{-2} + 14H^+ + 6e^-$ …………………(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 (ii), and then add to Eq. (ii).

$3S_2O_8^{-2} + 6e^- \rightarrow 6SO_4^{-2}$

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

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

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

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

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

d) $\boldsymbol{Mn0_4^- + C_2O_4^{-2} \rightarrow Mn^{+2} + CO_2}$

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

Reduction half-reaction: $\boldsymbol{Mn0_4^- \rightarrow Mn^{+2}}$

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

Step 2: Balance each half-reaction separately.

Reduction half-reaction:

Balance Mass:

Mn atoms are already balanced.

$$Mn0_4^- \rightarrow Mn^{+2}$$

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

$$Mn0_4^- + 8H^+ \rightarrow Mn^{+2} + 4H_2O$$

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.

$$Mn0_4^- + 8H^+ + 5e^- \rightarrow Mn^{+2} + 4H_2O \qquad \text{..............................(i)}$$

Oxidation half-reaction:

Balance Mass: Multiply CO_2 with 2 on the right side to balance C atoms.

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

O atoms are also balanced.

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

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.

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

Step 3: Balance the loss and gain of electrons in both half reactions and add them. The reduction half reaction has $5e^-$ while the oxidation half-reaction has $2e^-$. Thus, multiply Eq (i) by (ii), Eq (ii) by (v) and then add the resulting equations.

$$2Mn0_4^- + 16H^+ + 10e^- \rightarrow 2Mn^{+2} + 8H_2O$$

$$5C_2O_4^{-2} \rightarrow 10CO_2 + 10e^-$$

$$2Mn0_4^- + 16H^+ + 5C_2O_4^{-2} \rightarrow 2Mn^{+2} + 8H_2O + 10CO_2 + 10e^-$$

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

Duplications are $10e^-$. Cancel this out from both the sides.

$$2Mn0_4^- + 16H^+ + 5C_2O_4^{-2} \rightarrow 2Mn^{+2} + 8H_2O + 10CO_2$$

Q.10 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.

Q.11. 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.

Q.12. Explain Galvanic cell in detail. Draw a labeled diagram of a cell containing Zn-Cu electrodes.

The Galvanic Cell:

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

The electrodes at which oxidation occurs is called the anode. Whereas, the electrode at which 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:

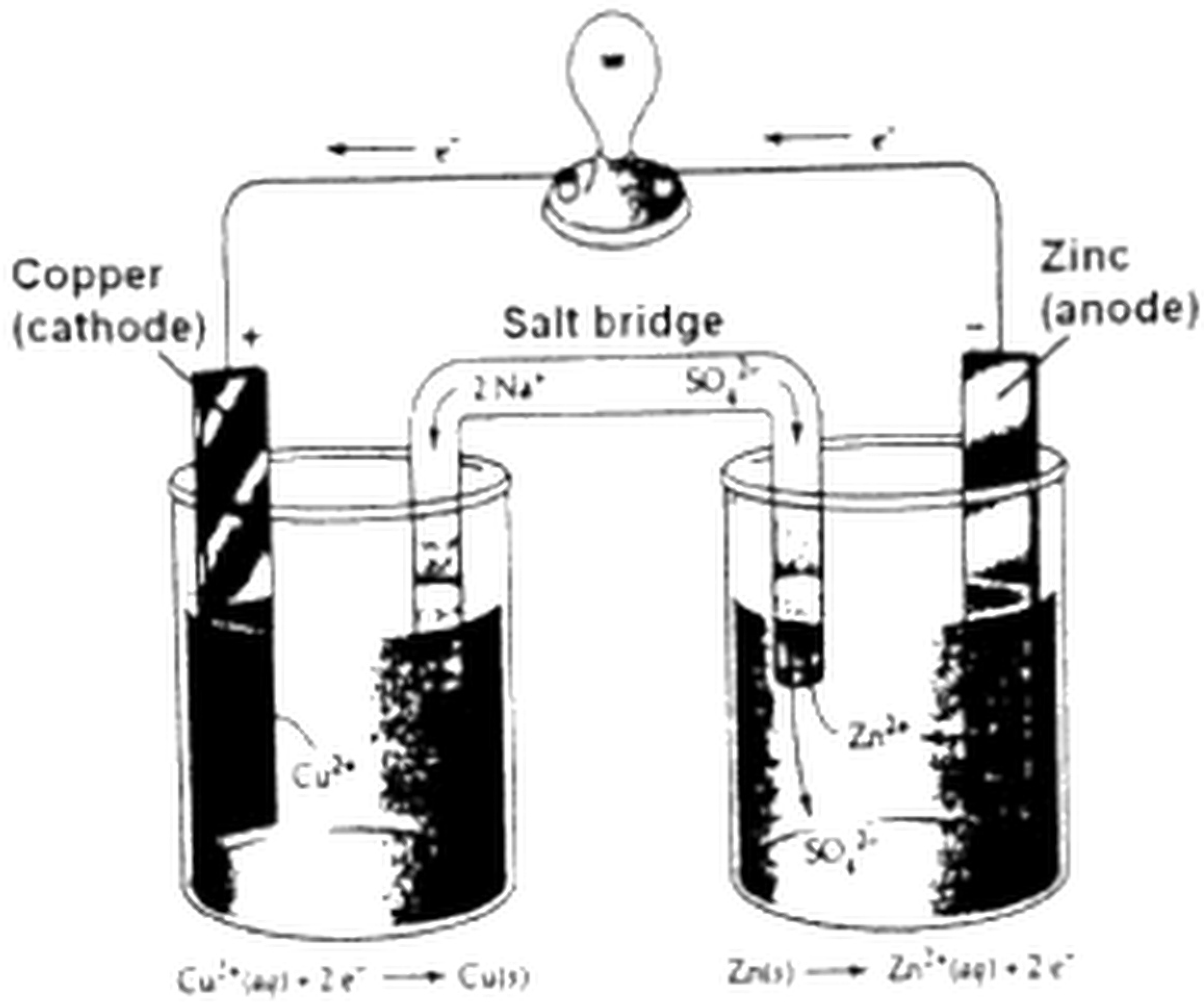
$$Zn(s) + Cu^{+2}(aq) \rightarrow Zn^{+2}(aq) + Cu(s)$$

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: $Zn(s) \rightarrow Zn^{+2}(aq) + 2e^-$ **(oxidation half reaction)**

At cathode: $Cu^{+2}(aq) + 2e^- \rightarrow Cu(s)$ **(reduction half reaction)**

Overall cell reaction: $Zn(s) + Cu^{+2}(aq) \rightarrow Zn^{+2}(aq) + Cu(s)$

Q.13. 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 ($E°$ 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, the 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 includes 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 or 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 more positive Standard potential does under the standard conditions (1 mol dm^{-3} concentration at 25° C and 1 atm pressure), the reaction with more negative E° value occurs as oxidation (anode reaction). The reaction with more positive E° 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.

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

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

Measurement of electrode potential:

The potential of galvanic cell can be measured with the voltmeter. But a single half-cell potential of 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 designing 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 deduction relative to the standard hydrogen electrode.

Explanation:

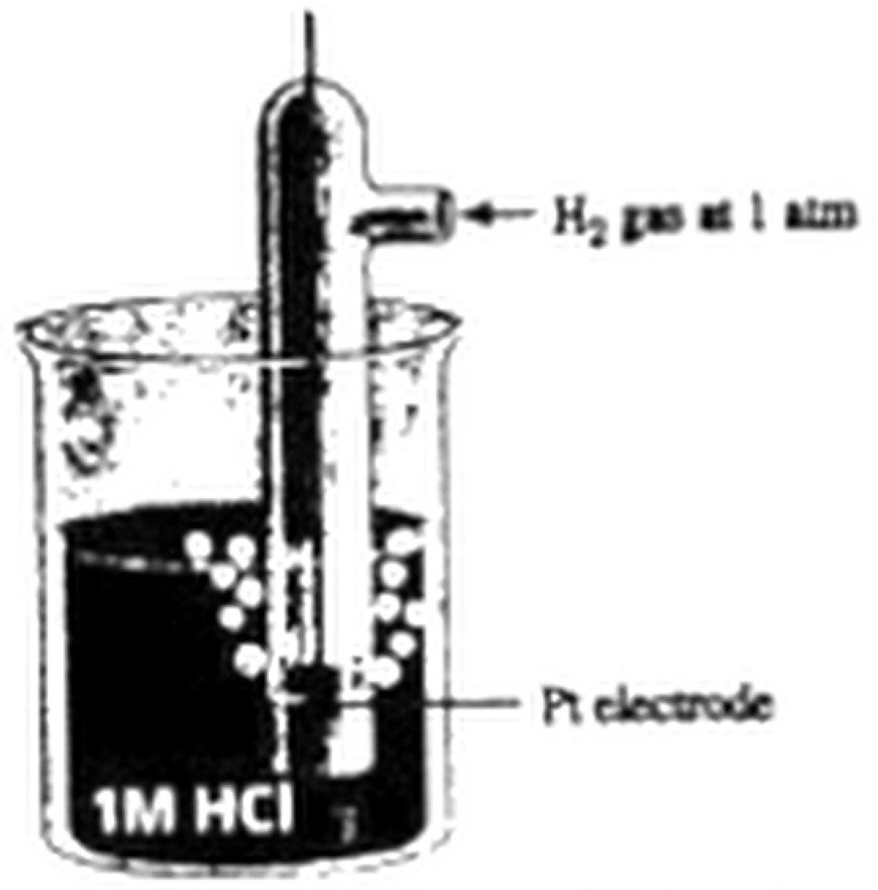
It is potential difference developed when an electrode of an element is placed in a solution containing ions of that element when all the components are in their standard state i.e. 1 atm of a gases, 1M for solutions, pure solid for electrode and at 25° C.

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

Ans: Standard Hydrogen Electrode:

Construction:

A standard hydrogen electrode (SHE) consists of a Platinum coated with finely divided Platinum surrounded by hydrogen gas at 1 atm 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):

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

Oxidation: (act as anode):

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

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

Reduction Half-reaction	E0 (Volts)
$Li^{+} + e^{-} \rightleftharpoons Li$	-3.05
$K^{+} + e^{-} \rightleftharpoons K$	-2392
$Ba^{2+} + 2e^{-} \rightleftharpoons Ba$	-2.90
$Ca^{2+} + e^{-} \rightleftharpoons Na$	-2.76
$Na^{2+} + 2e^{-} \rightleftharpoons Na$	-2.71
$Mg^{2+} + 2e^{-} \rightleftharpoons Mn$	-2.38
$Al^{3+} + 3e^{-} \rightleftharpoons Al$	-1.67
$Mn^{2+} + 2e^{-} \rightleftharpoons Mn$	-1.03
$2H_2O + 2e^{-} \rightleftharpoons H_2:.2OH^{-}$	-0.83
$Zn^{2+} + 2e^{-} \rightleftharpoons Zn$	-0.76
$Cr^{3+} + 3e^{-} \rightleftharpoons Cr$	-0.74
$Fe^{2+} + 2e^{-} \rightleftharpoons Fe$	-0.44
$PbSO_4 + 2e^{-} \rightleftharpoons Pb + SO_4^{2-}$	-0.36
$Ni^{2+} + 2e^{-} \rightleftharpoons Ni$	-0.25
$Sn^{2+} + 2e^{-} \rightleftharpoons Sn$	-0.14
$Pb^{2+} + 2e^{-} \rightleftharpoons Pb$	-0.13
$Fe^{3+} + 3e^{-} \rightleftharpoons Fe$	-0.04
$2H^{+} + 2e^{-} \rightleftharpoons H_2$	0.00
$AgCl + e^{-} \rightleftharpoons Ag + Cl^{-}$	+0.22
$Hg_2Cl_2 + 2e^{-} \rightleftharpoons 2Hg + 2Cl^{-}$	+0.27
$Cu^{2+} + 2e^{-} \rightleftharpoons Cu$	+0.34
$Cu^{+} + 1e^{-} \rightleftharpoons Cu$	+0.52
$I_{2(aq)} + 2e^{-} \rightleftharpoons 2I^{-}$	+0.54
$Fe^{3+} + e^{-} \rightleftharpoons Fe^{2+}$	+0.77
$Ag^{+} + e^{-} \rightleftharpoons Ag$	+0.80

$Br_{(aq)} + 2e^- \rightleftharpoons 2Br^-$	+1.09
$O_2 + 4H^+ + 4e^- \rightleftharpoons 2H_2O$	+1.23
$MnO_2 + 4H^+ + 2e^- \rightleftharpoons Mn^{2+} + 2H_2O$	+1.28
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightleftharpoons 2Cr^{3+} + 7H_2O$	+1.33
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-$	+1.36
$2ClO_3 + 12H^+ + 5e^- \rightleftharpoons Cl_2 + 6H_2O$	+1.47
$8H^+ + MnO_4 + 5e^- \rightleftharpoons Mn^{2+} + 4H_2O$	+1.49
$PbO_2 + SO_4^{-2} + 4H^+ + 4e^- \rightleftharpoons PbSO_4 + 2H_2O$	+1.69
$H_2O_2 + 2H^+ + 2e^- \rightleftharpoons 2H_2O$	+1.7
$S_2O_3^{-2} + 2e^- \rightleftharpoons 2SO_4^{-2}$	+2.00
$F_2 + 2e^- \rightleftharpoons 2F^-$	+2.87

Table: Reduction potentials of some elements, ions and compounds

Q 16. 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 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° 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 the 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.

$Zn (s) \longrightarrow Zn^{+2} (aq) + 2\bar{e}$ (Oxidation half reaction)

$Cu^{+2} (aq) + 2\bar{e} \longrightarrow Cu (s)$ (Reduction half reaction)

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

$$E^\circ_{cel} = E^\circ_{cathode} - E^\circ_{anode}$$

$$E^\circ_{cel} = E^\circ_{Cu} - E^\circ_{Zn}$$

$$E^\circ_{cel} = +0.34 - (-0.76)$$

$$E^\circ_{cel} = +1.10V$$

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:

$Ni^{+2} (aq) + 2\bar{e} \longrightarrow Ni (s)$ $E^\circ = -0.25V$

Mg^{+2} (aq) + $2\bar{e}$ $\longrightarrow$ Mg (s) $E^{\circ} = -2.38V$

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

Solution:

Data indicates that reduction potential of first reaction is greater than that of 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) + $2\bar{e}$ (Anode reaction)

Ni^{+2} (aq) + $2\bar{e}$ $\longrightarrow$ Ni (s) (Cathode reaction)

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

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

$$E^{\circ}_{cel} = E^{\circ}_{cathode} - E^{\circ}_{anode}$$

$$E^{\circ}_{cel} = E^{\circ}_{Ni} - E^{\circ}_{Mg}$$

$$= -0.25 - (-2.38) = 2.13\ V$$

SELF-CHECK EXERCISE 12.7

The standard reduction potential for the following half reactions are:

Cu^{+2} (aq) + 2 e^- $\longrightarrow$ Cu (s) $E^{\circ} = +0.34V$

Cl_2 (g) + 2 e^- $\longrightarrow$ 2 Cl (aq) $E^{\circ} = +1.36V$

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

Solution:

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

Anode Reaction: $Cu\ (s) \longrightarrow Cu^{+2}(aq) + 2\ e^-$ (Oxidation half reaction)

Cathode Reaction: $Cl_2\ (g) + 2\ e^- \longrightarrow 2Cl^-(aq)$ (Reduction half reaction)

Cell reaction: $Cu\ (s) + Ni^{+2} \longrightarrow Cu^{+2}\ (aq) + 2Cl^-\ (aq)$ (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

$$E^\circ_{cel} = E^\circ_{cathode} - E^\circ_{anode}$$

$$E^\circ_{cel} = E^\circ_{Cl2} - E^\circ_{Cu}$$

$$= +1.36 - (+0.34) = +1.02\ V$$

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° values of two half-cell reactions.

Condition:

i. The positive value indicates 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 $E^\circ_{Sn} = -0.14V$, $E^\circ_{Fe} = -0.44V$

Solution:

Oxidation

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

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.

$Sn \qquad Sn^{+2} + 2\bar{e}$ (Oxidation half reaction)

$Fe^{+2} + 2\bar{e} \qquad Fe$ (Reduction half reaction)

$E^\circ_{cel} = E^\circ_{cathode} - E^\circ_{anode}$

$E^\circ_{cel} = -0.44V - (-0.14V) = -0.30$

As E°_{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. $E^\circ_{Cu} = \mathbf{0.34V}$

$E^\circ_{Fe} = -0.448V$

$$E^\circ_{cel} = E^\circ_{cathode} - E^\circ_{anode}$$

$$E^\circ_{cel} = E^\circ_{Cu} - E^\circ_{Fe}$$

$$E^\circ_{cel} = +0.34V - (-0.44V) = +0.78V$$

As E°_{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. $E^\circ_{I} = \mathbf{0.54V}$

$E^\circ_{Br} = 1.09V$

$$E^\circ_{cel} = E^\circ_{cathode} - E^\circ_{anode}$$

$$E^\circ_{cel} = E^\circ_{I} - E^\circ_{Br}$$

$$E^\circ_{cel} = +0.54V - (+1.09V) = -0.55V$$

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

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 used for wrapping candies biscuits etc..) in such a way that the foil presses against dental filling, will experience a momentarily sharp pain.

This is because an electrochemical 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 feeling short circuits the cell. This causes the weak current to flow between the electrodes. This current simulates 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 reaction 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 the electrode.

Note: The given reduction potential values related to standard conditions only i.e. 1 M solution of ions, 25° C (298 k) and 1 atm pressure. Changes in condition will after these values. Such a list of arrangement of elements in 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 reactions 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 a standard reduction potential reaction o 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 — Very Active metals;
Ba — react with cold water with the liberation of hydrogen gas;
Sr — (K and Na react violently with water).
Ca — they also react violently with acids
Na

Mg
Al
Mn — Metals of intermediate activity:
Zn — React with steam or with acids such as HCl with liberation of H_2
Cr
Fe
Cd

Co
Ni — Moderately active metals
Sn — React slowly with HCl
Pb — 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 of activity series of common metals are as follows:

1. Metals high on the list transfer electrons to metal questions law and 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 ion pallets 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 reacts 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 the hydrogen in the activity series do not react with 1M HCl. These metals are unable to reduce H^+ ions n 1M HCl solution. However, their 1M aqueous ions can be reduced by hydrogen gas at 1 atm. These metals are called noble metals.

The moderately noble metals such as Cu, Ag And Hg react only with oxidizing acids such as nitric and perchloric acids. The very noble metals do not 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 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^{*}.

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

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

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

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 occur:

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

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

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

Solution:

a) Iron is ranked higher in activity series then hydrogen. Thus, it can displace hydrogen from HCL and the following reaction will occur.

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

b) Zinc is ranked higher in activity series then 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 then silver. Thus, it can displace silver from silver nitrate and the following reaction will occur.

$$Ni + 2AgNO_3 \longrightarrow Ni\ (NO_3) + 2Ag$$

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, platting of noble metal on jewellery and instruments.

Electrolysis:

The chemical process used in electrolytic cells is called electrolysis.

Electrolysis is a production of a chemical reaction by means of an electric Current. The

Apparatus used of electrolysis consists of an electrolytic cell.

Construction:

Electrolytic cells 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 carried from the battery through the cause by means of electrons

(metallic conduction).

Working:

i. within the cell the current is carried by the actions and cations of the electrolyte (Electrolytic condition).

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

iii. At each electrode a chemical reaction takes pace 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 return to the battery do the connecting wires.

vi. Thus oxidation reduction reactions occur at the electrodes.

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

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

Note: The changes which occurs at the electrodes, depends upon the relative oxidation-reduction tendencies of a substance 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 change 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 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 describe 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:

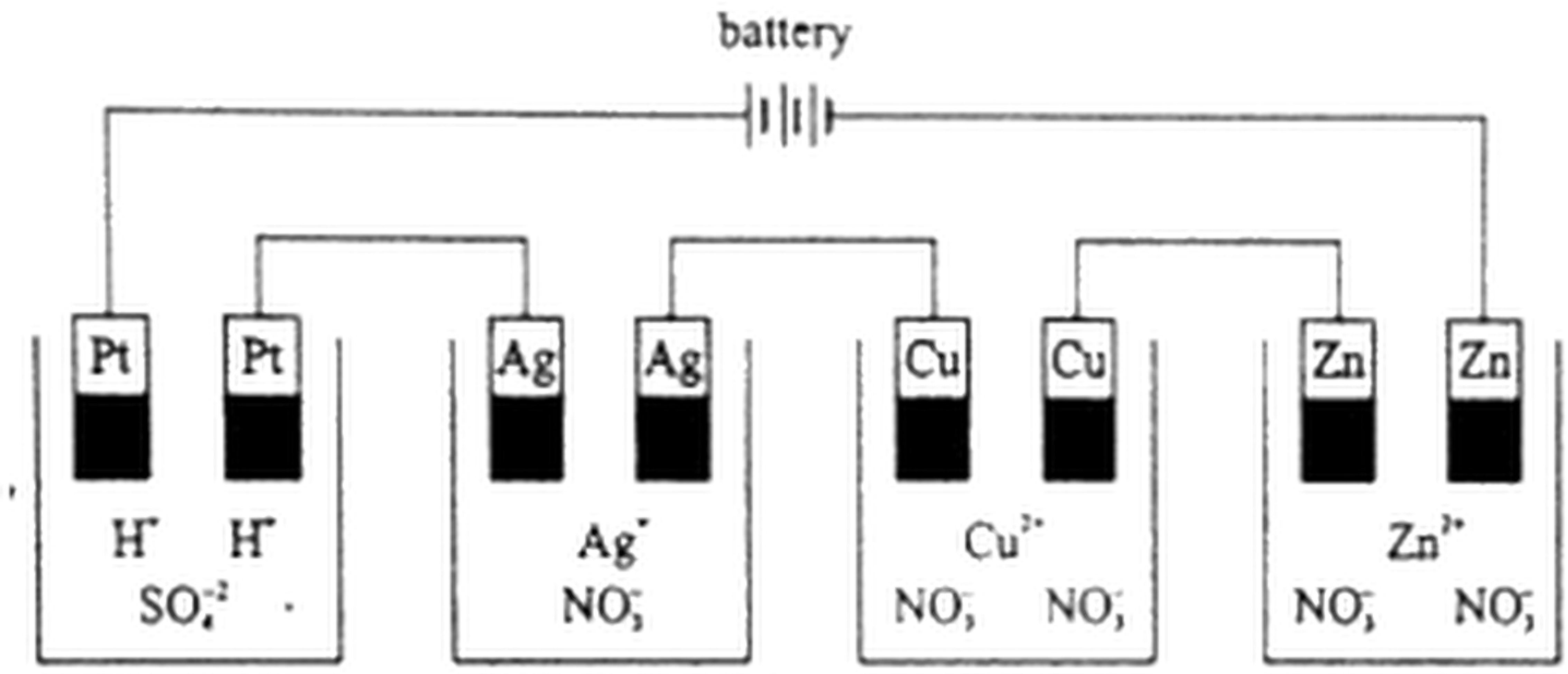
The first law of Faraday states that 'the amount of chemical reaction taking place at an electrode is directly proportional to the quantity of charge that flows through the electrode during the process.'

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 parents 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 proceeds simultaneously and for 7.999g of O_2 from H_2SO_4, 1.008g of H_2, 107.9g of Ag. 31.777g 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 change on one mole of electrons. This quantity of charge is referred as one Faraday. Thus, the quantity of the change that occur 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 $= 0.01 \times 1 \times 60 \times 60 = 36\ C.$

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

In molten zinc chloride the cathode reaction is

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

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

Thus,

3.7×10^{-4} Faraday $\times$ 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 $\times 63.37g = 0.012g$ 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 of cathode is the reduction of Au (III) to Au metal

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

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

Moles of Au = $\frac{1.314 gAu}{197g\ /mole\ of\ Au}$ $= 6.67 \times 10^{-3}$

i) charge = 6.67×10^{-3} mole of Au $\times \frac{3\ Faraday}{mole\ of\ Au}$

= 2×10^{-2} Faraday

ii) **current** **=** $\frac{charge}{Time\ (s)}$

Time = 10 mins = $10 \times 60 = 600s$

Current = $\frac{(2 \times 10^{-2} F)(96500C\ /F}{600s} = 3.22A$

iii) The reaction of 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 = ½ moles of Cl_2 was produced.

For 2×10^{-2} Faraday of electricity = ½ × 2×10^{-2} moles of Cl_2 was produced.

= 1×10^{-2} moles 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:

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

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 $= R = 0.0821\ atm\ dm^{-3}\ Mole\ K^{-1}$

Time: $= 10\ hours = 10 \times 60 \times 60 = 36000s$

Current $= 15A$

$\boldsymbol{Current} = \frac{charge}{Time(s)}$

$Charge = Current \times Time(s)$

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

Mass of Al:

In the electrolysis of Al_2O_3. The cathode reaction is

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

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

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

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

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

Volume of O_2:

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

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

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 = 3600s

Current = 15A

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

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

$$= 15A \times 3600(s) = 54000\ C$$

As 96500 C = 1 F

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

a) NaCl:

In the electrolysis of molten $NaCl$. The cathode reaction is

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

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

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

1F charge produces Na = 1 mole

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

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

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

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

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

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

1F = 1 mole of electron

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

2F charge produces Ca = 1 mole

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

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

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

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

Thus, NaCl will produce more mass of Na

Q.24 What are batteries? Also give their applications.

Ans: Batteries:

A battery is a Galvanic cell or group of Galvanic cell connected in series. Batteries are a 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.

Q.25 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).

Q.26 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……………………………………………………………… 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 $NH_4Cl \text{ and } ZnCl_2$

Reactions: Following reactions take place in it.

At Anode: $Zn \rightarrow Zn^{+2} + 2e^-$

At Cathode: $2NH_4^+ + 2MnO_2 + 2e^- \rightarrow Mn_2O_3 + 2NH_3 + H_2O$

This cell produces a potential of 1.5V.

Types:

(i) Alkaline dry cell:

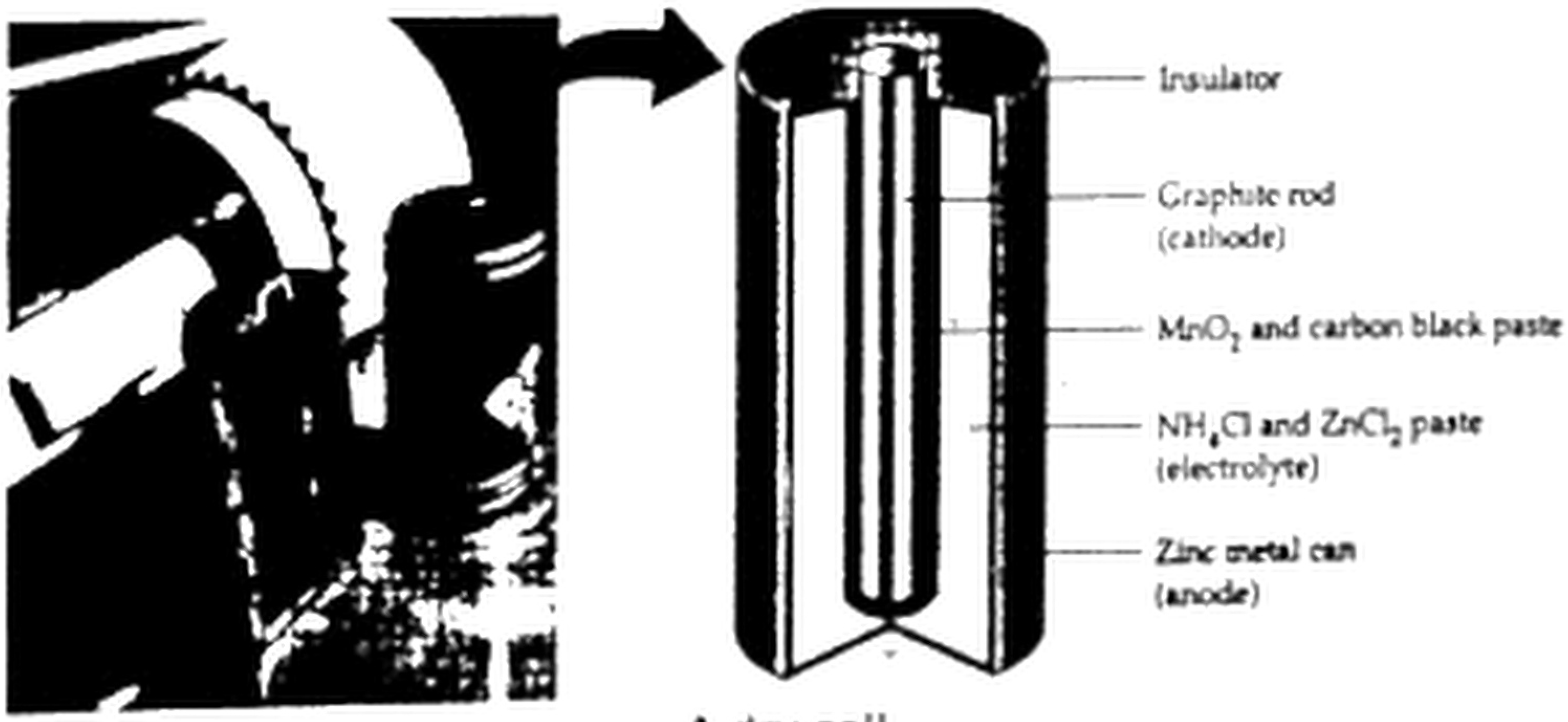
Construction:

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

Reactions: Following reactions take place in it.

At Anode: $Zn + 2OH \rightarrow Zn(OH)_2 + 2e^-$

At Cathode: $2MnO_2 + H_2O + 2e^- \rightarrow Mn_2O_3 + 2OH^-$



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-cadmium dry cell:

An important type of dry cell is the nickel-cadmium battery.

Construction:

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

Reactions: Following reactions occur in it.

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

At Cathode: $NiO_2(s) + 2H_2O(l) + 2e^- \rightarrow Ni(O)_2(aq) + 2OH^-(aq)$

Recharging:

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

Q.27. 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.

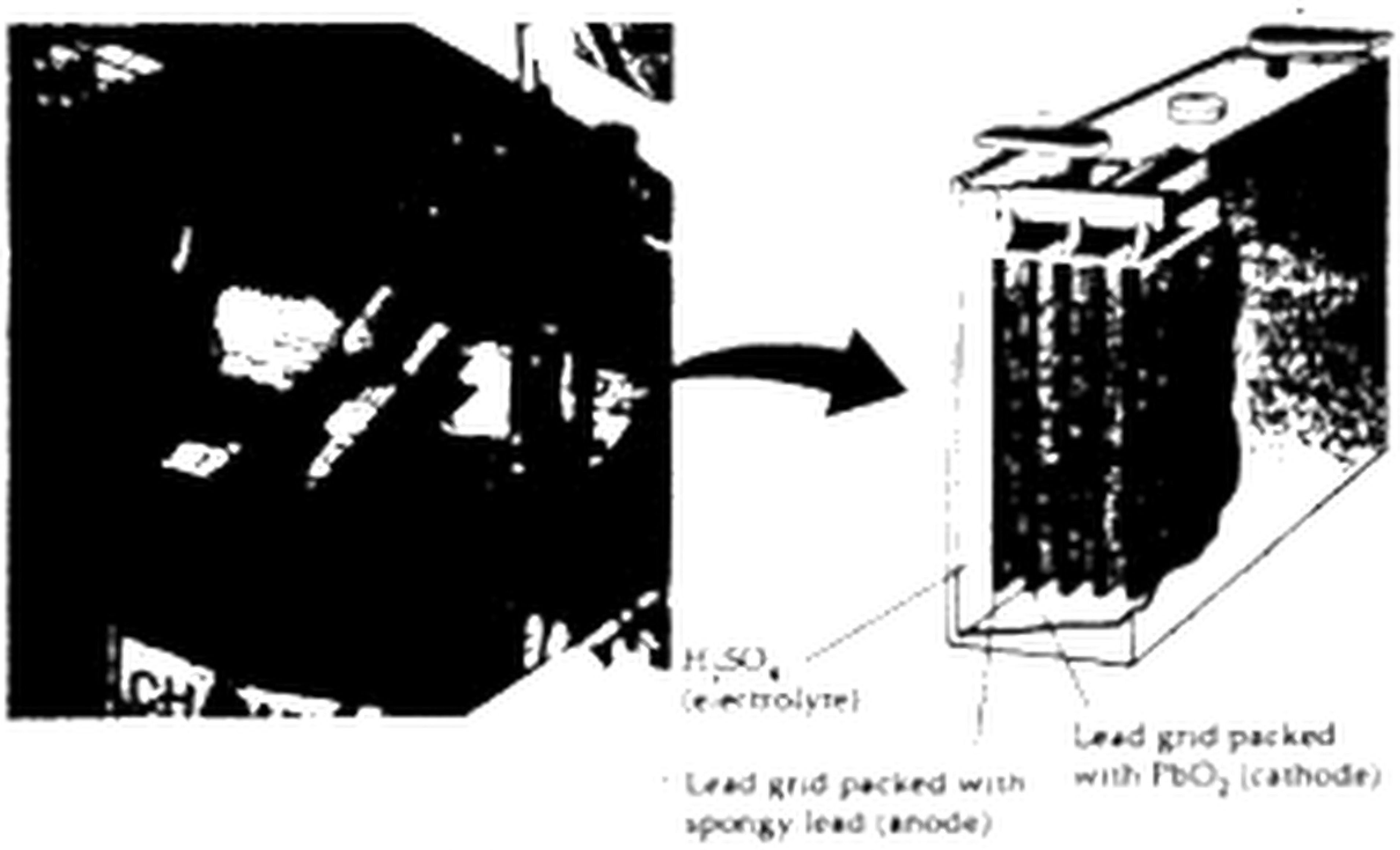
ii. The cathode is lead impregnated with PbO_2 which also becomes coated with $PbSO_4$

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: $Pb(s) + SO_4^{-2}(aq) \rightarrow PbSO_4(s) + 2e^-$

At Cathode: $PbO_2(s) + SO_4^{-2}(aq) + 4H^+(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$

Net reaction during discharge of battery is

$$Pb_{(s)} + PbO_{2(s)} + 2H_2SO_{4(aq)} \rightarrow 2PbSO_{4(s)} + 2H_2O_{(l)}$$

Thus during discharge H_2SO_4 is used up and its density decreases. When both 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 reaction during recharging:

The reactions of recharging of battery are as follows.

At cathode:

$$PbSO_{4(s)} + 2e^- \rightarrow Pb_{(s)} + SO_4^{-2}{}_{(aq)}$$

At anode: $\boldsymbol{PbSO_{4(s)} + 2H_2O_{(l)} \rightarrow PbO_{2(s)} + SO_4^{-2}{}_{(aq)} + 4H^+{}_{(aq)} + 2e^-}$

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

After recharging H_2SO_4 solution is concentrated again bringing density to its initial value of 1.25 g cm.

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.

Q.28. What are fuel cells? How they work? Also write the redox reactions.

Ans: Fuel Cells:

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.

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

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.

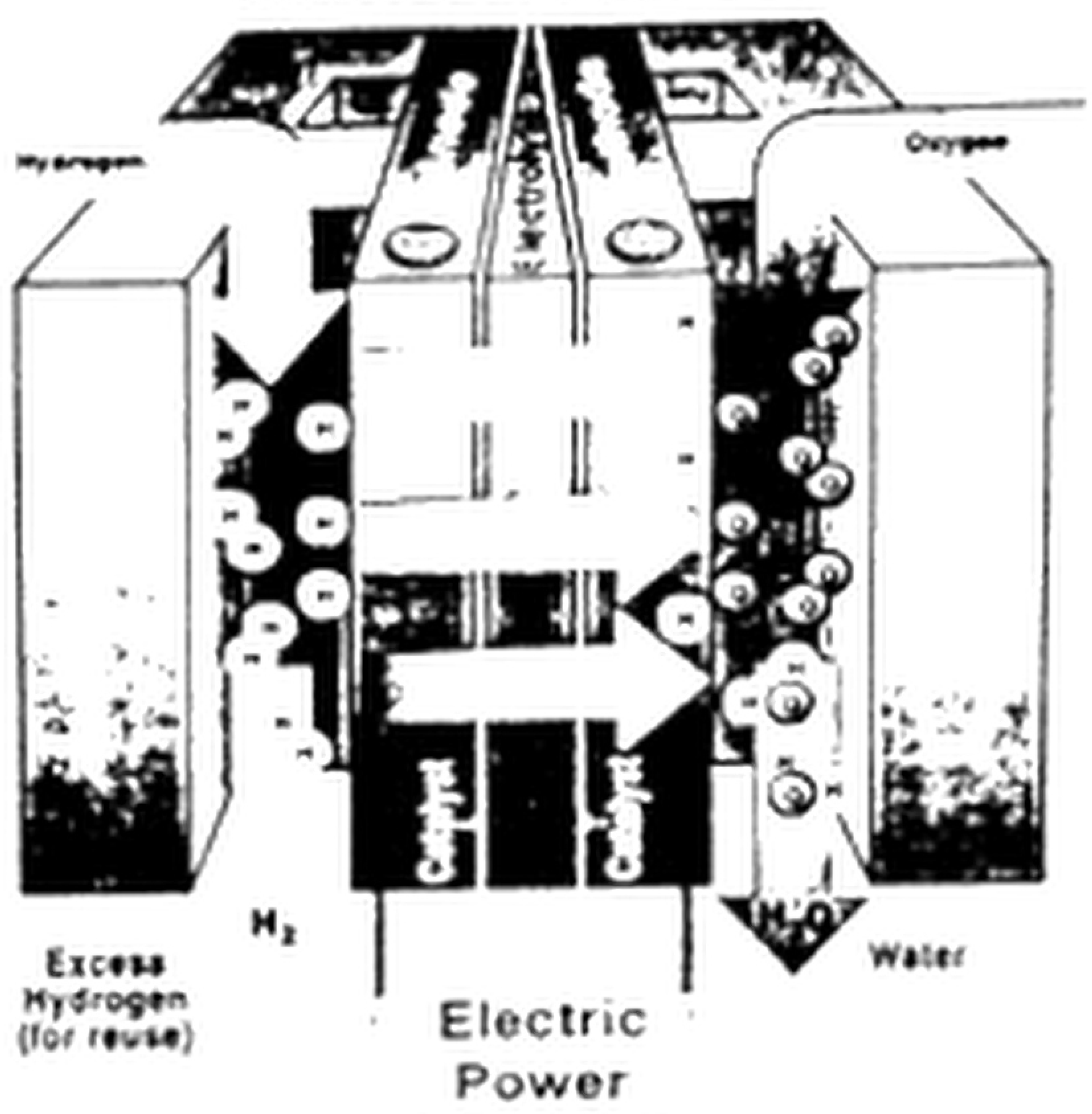
Construction:

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.



A Fuel Cell

Redox reaction:

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

Anode reaction: = $2H_2(g) + 4OH^-(aq) \rightarrow 4H_2O(l) + 4e^-$

Cathode reaction: = $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$

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

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.

Q.29. Define solar cell. How can we get electrical energy from it?

Ans: Solar cell:

Devices that convert solar energy directly into electrical 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 close 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.

Q.30. 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

$$Fe^{+2} + 2e^- \rightarrow Fe \qquad E^\circ = -0.44V$$

Reverse of this reaction

i.e. $Fe \rightarrow Fe^{+2} + 2e^-$

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 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:

$$O_2(g) + H^+(aq) + e^- \rightarrow 2H_2O(l) \qquad E^\circ = +1.23V$$

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

$$Fe \rightarrow Fe^{+2} + e^{-} \quad \text{(oxidation reaction)}$$

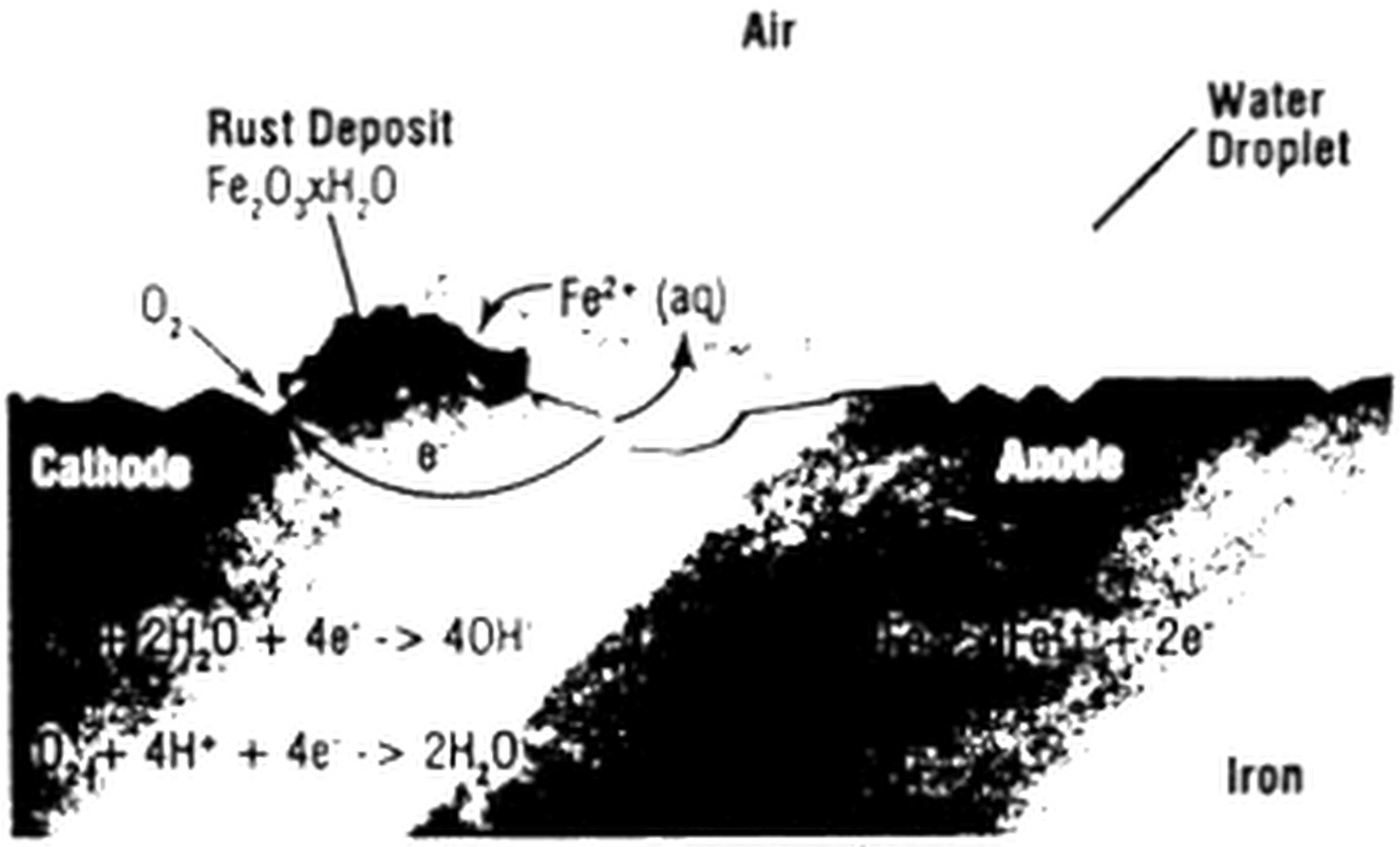
The $E°$ cell of the combined half-reaction is

$$E° \, cell = E°_{cathode} - E°_{anode}$$

$$E° \, cell = +1.23 - (-0.44)$$

$$E° \, cell = +1.67V$$

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 metal can be prevented 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.

Q.31. 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.

$$Fe^{+2} + 2e^- \rightleftharpoons Fe \quad E^\circ = -0.44V$$

$$Zn^{+2} + 2e^- \rightleftharpoons Zn \quad E^\circ = -0.76V$$

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.

Q.32. 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).

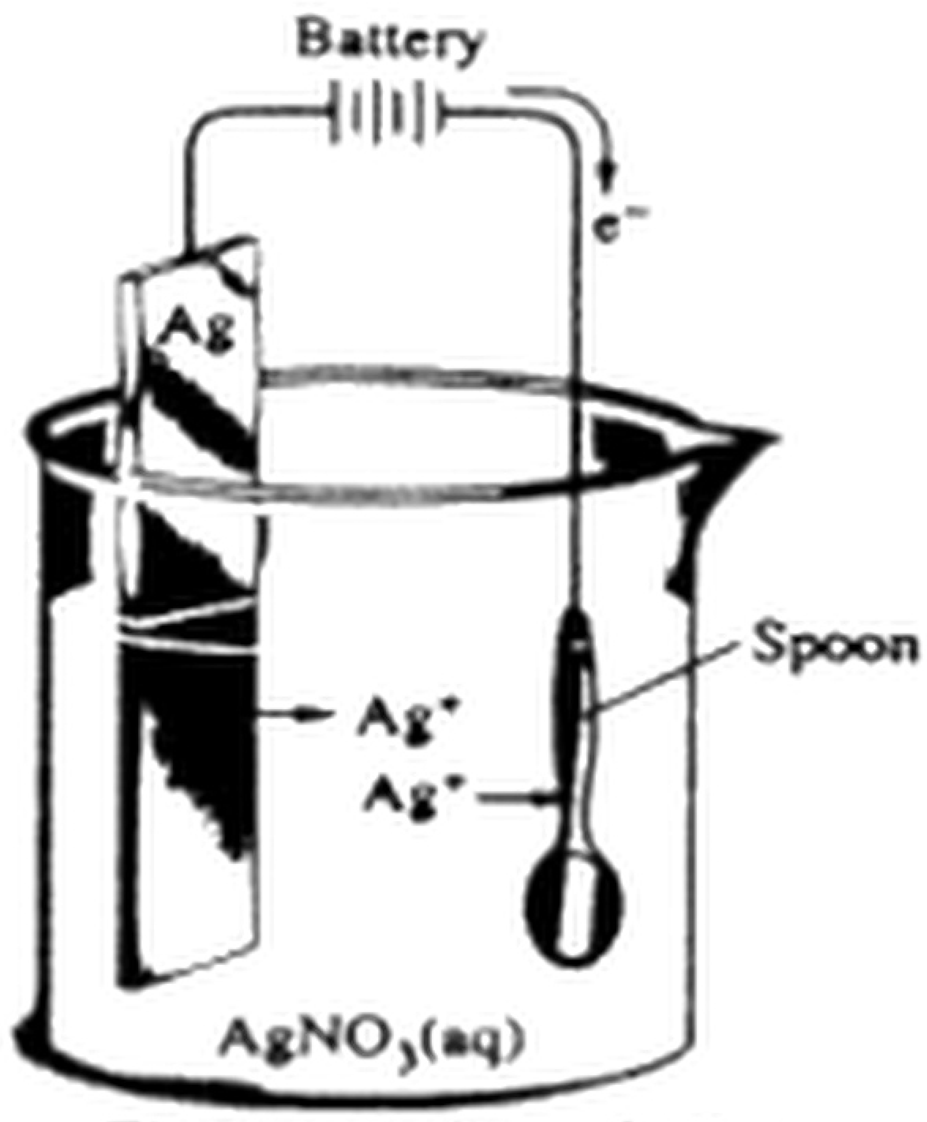
Example:

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

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

Anode reaction: $Ag\ (s) \rightarrow Ag^{+}(aq) + e$

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



Electroplating of sliver

Use:

Steel objects are often protected from corrosion by electroplating with Aluminum or Chromium. These metals forma 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 a 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 reaction take place at separate electrodes and electron flow through the external circuit. These separate parts of the Galvanic 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.