# 7 Redox Reactions


*Breath alcohol analyzers are brilliant pieces of technology that rely on the redox reaction. The driver's lungs' air is drawn into the test vial, which contains a reddish-orange mixture of potassium dichromate, sulphuric acid, and silver nitrate. This mixture, together with the ethanol from one's breath, constitutes a redox reaction system. The demarcation on the knob aids police in determining whether or not a person is intoxicated.*

## Topic Notes

- *Basic Concepts of Redox Reactions*

# BASIC CONCEPTS OF REDOX REACTIONS

# 1

## TOPIC 1

## CLASSICAL IDEA OF REDOX REACTION, OXIDATION AND REDUCTION

### Redox Reaction

Redox reaction is defined as the reaction in which oxidation and reduction take place simultaneously. All these reactions are accompanied by energy transfer in the form of heat, light or electricity.

### Oxidation

It is defined as the addition of oxygen or an electronegative element to a substance or the removal of hydrogen/electropositive element from a substance.

For example,

$$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$$

$$S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$$

In the above two reactions, the magnesium and sulphur undergo oxidation to form magnesium oxide and sulphur dioxide. One more example of oxidation is where removal of the electropositive element is taking place:

$$2H_2S_{(g)} + O_{2(g)} \rightarrow 2S_{(s)} + 2H_2O_{(l)}$$

In the given reaction, the hydrogen is removed from hydrogen sulphide and combined with oxygen to form water.

### Reduction

It is defined as the addition of hydrogen/electropositive element to a substance or the removal of the oxygen/electronegative element from a substance.

For example,

$$2HgO_{(s)} \xrightarrow{\Delta} 2Hg_{(l)} + O_{2(g)}$$

$$2FeCl_{3(aq)} + H_{2(g)} \xrightarrow{\Delta} 2FeCl_{2(aq)} + 2HCl_{(aq)}$$

In the first reaction, the oxygen is removed from the mercuric oxide while in the second reaction, the removal of an electronegative element that is chlorine from the ferric chloride takes place.

## TOPIC 2

## REDOX REACTION IN TERMS OF ELECTRON TRANSFER

According to the electronic concept, every redox reaction in terms of electron transfer contains two steps which are known as half-reactions.

### Oxidation Reaction

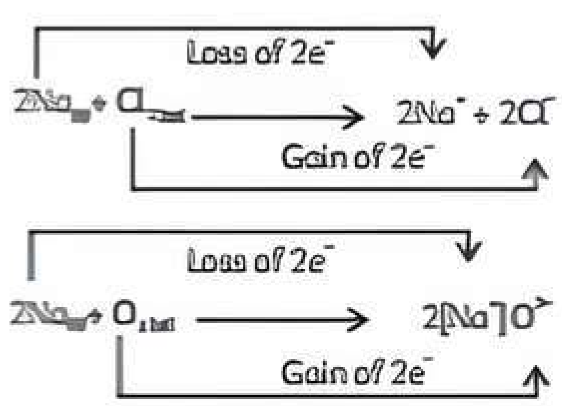
According to this concept, a half-reaction that involves the loss of the electrons is called an oxidation reaction.

### Reduction Reaction

According to this concept, a half-reaction that involves the gain of the electrons is called a reduction reaction.

### Redox Reaction

The reaction in which both half-reactions, i.e. oxidation reactions (loss of electrons) and reduction reactions (gain of electrons) occurring simultaneously is known as a redox reaction. For example:




In these reactions, the sodium loses two electrons, thus the half-reaction is known as an oxidation reaction and chlorine and oxygen gain two electrons, thus this half-reaction is known as a reduction reaction. As both the reactions are occurring simultaneously, thus the reaction is known as the redox reaction.

## Competitive Electron Transfer Reactions

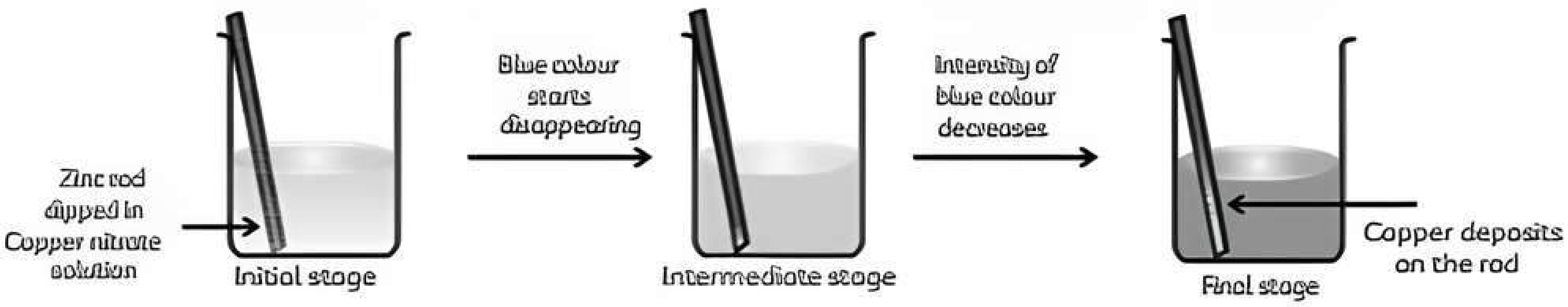
We understand this concept through an experiment:

(1) Place a strip of metallic zinc in an aqueous solution of copper nitrate. After one hour following changes will be noticed.

(2) Strip becomes coated with reddish metallic copper.

(3) Blue colour of the solution disappears.

(4) If hydrogen sulphide gas is passed through the solution appearance of white ZnS can be seen in making the solution alkaline with ammonia.




*Redox reaction between zinc and an aqueous solution of copper nitrate*

The reaction that takes place in the aqueous solution can be shown as:

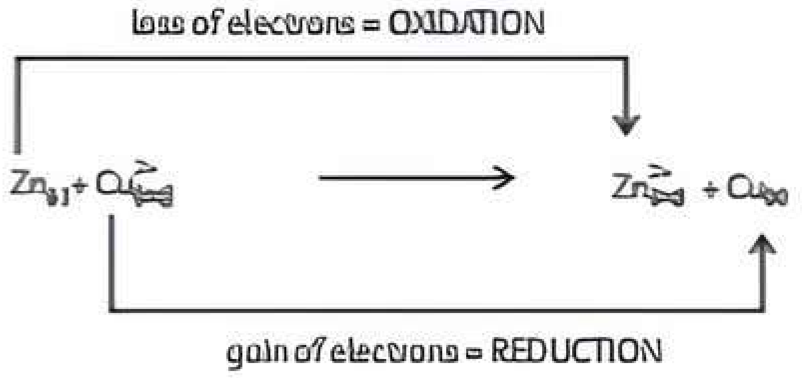
Oxidation half-reaction: $Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2e^-$

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

Overall redox reaction:

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

This is how the reaction takes place:




## Oxidising and Reducing Agent

**Oxidising agent**

The species in a reaction which gains electrons is called an oxidising agent. They are also known as oxidants.

**Reducing agent**

The species in a reaction which loses electrons is called a reducing agent. They are also known as reductants.

**Example 1.1:** **Identify the reducing and oxidising agents in the balanced redox reaction:**

$Cl_{2(aq)} + 2Br^-{}_{(aq)} \rightarrow 2Cl^-{}_{(aq)} + Br_{2(aq)}$

**Ans.** Oxidation half-reaction:

$$2Br^-{}_{(aq)} \rightarrow Br_{2(aq)}$$

Oxidation half-reaction:

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

Here, $Br^-$ loses an electron; it is oxidised from $Br^-$ to $Br_2$; thus, $Br^-$ is a reducing agent. $Cl_2$ gains an electron; it is reduced to $Cl^-$; thus, $Cl_2$ is an oxidising agent.

# TOPIC 3

# OXIDATION NUMBER

It is the oxidation state of an element in a compound which is the charge assigned to an atom of a compound is equal to the number of electrons in the valence shell of an atom that is gained or lost completely or to a large by that atom while forming a bond in a compound.

**Note:** Oxidation number denotes the oxidation state of an element in a compound.

**Important**

➥ *Stock notation is the oxidation number expressed by putting a roman numeral representing the oxidation number in parenthesis after the symbol of the metal in the molecular formula.*

(1) The oxidation number of an element in its elementary form is zero. For example, $H_2$, $O_2$ and $N_2$. They have an oxidation number equal to zero.

(2) In a single monoatomic ion, the oxidation number is equal to the charge in the ion. For example, $Na^+$ ion has an oxidation number of +1 and $Mg^{2+}$ ion has +2.

(3) Oxygen has an oxidation number −2 in its compounds.

(4) In non-metallic compounds of hydrogen-like HCl, $H_2S$ and $H_2O$ oxidation number of hydrogen is +1 but in metal hydrides, oxidation number of hydrogen is −1 [LiH, NaH, $CaH_2$].

(5) In compounds formed by the union of metals with non-metals, the metals have positive oxidation numbers while non-metals have negative oxidation numbers. For example, In NaCl, Na has a +1 oxidation number while Cl has –1.

(6) In a compound, there are two non-metallic atoms the atoms with high electronegativity are assigned a negative oxidation number while other atoms have a positive oxidation number.

(7) In neutral compounds, the algebraic sum of the oxidation number of all atoms in a compound is equal to zero.

(8) In a polyatomic ion, the sum of the oxidation no. of all the atoms in the ion is equal to the net charge on the ion.

For example, in $(CO_3)^{2-}$, the sum of the oxidation number of a carbon atom and three oxygen atoms is equal to –2.

### Important

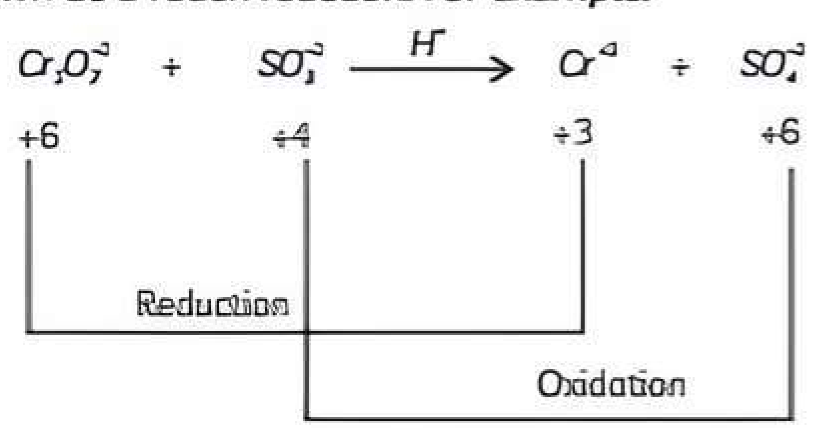
➥ *There are some exceptions where the oxidation state of oxygen is +2. In compounds such as peroxides, i.e., $Na_2O_2$ and $H_2O_2$, the oxidation number of oxygen is –1.*

*Oxidation number of oxygen = +2 in $OF_2$*
*Oxidation number of oxygen = +1 in $O_2F_2$*

**Oxidation:** The process in which the oxidation number of an element increases is called oxidation.

**Reduction:** The process in which the oxidation number of an element decreases is called reduction.

**Redox reaction:** The reaction which involves a change in the oxidation number of the interacting species is known as a redox reaction. For example:




**Oxidising agent:** A reagent which can increase the oxidation number of an element in the given substance. These reagents are also known as oxidants.

**Reducing agent:** A reagent which lowers the oxidation number of an element in a given substance. These reagents are called as reductants.

### Important

➥ *Fractional Oxidation Number: Elements as such do not have any fractional oxidation numbers. When the same element is involved in different bonding in a species, their actual oxidation states are whole numbers but an average of these is fractional. Fractional O.N. of a particular element can be calculated only if we know about the structure of the compound or in which it is present. For example, In $C_3O_2$*

$$O = C = C = C = O$$

*The average O.N. of carbon atom =* $\dfrac{(2+2+0)}{3} = \dfrac{4}{3}$

**Example 1.2:** Case Based:

*The oxidation state of an atom (sometimes referred to as the oxidation number) in a chemical compound provides insight into the number of electrons lost and, therefore, describes the extent of oxidation of the atom. The oxidation state of an atom can be defined as the hypothetical charge that would be held by that atom if all of its bonds to other atoms were completely ionic in nature. The oxidation state of an atom is not regarded as the real charge of the atom. In a chemical reaction, if there is an increase in the oxidation state, then it is known as oxidation whereas if there is a decrease in oxidation state, it is known as reduction.*

(A) The reaction,

$Cl_{2(g)} + 2OH^-_{(aq)} \rightarrow ClO^-_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$

represents the process of bleaching. Identify and name the species that bleaches the substances due to their oxidising action.

(B) Identify the redox reactions out of the following reactions and identify the oxidising and reducing agents in them.

(i) $3HCl_{(aq)} + HNO_{3(aq)} \rightarrow Cl_{2(g)} + NOCl_{(g)} + 2H_2O_{(l)}$

(ii) $HgCl_{2(aq)}\ 2KI_{(aq)} \rightarrow HgI_{2(s)} + 2KCl_{(aq)}$

(iii) $Fe_2O_{3(s)} + 3CO_{(g)} \xrightarrow{\Delta} 2Fe_{(s)} + 3CO_{2(g)}$

(C) In which of the compounds sulphur will have an oxidation state of +6?

(a) $Na_2S$ (b) $Na_2S_4O_6$
(c) $Na_2SO_3$ (d) $Na_2SO_4$

(D) Assertion (A): In the reaction between potassium permanganate and potassium iodide, permanganate ions act as oxidising agents.

Reason (R): Oxidation state of manganese changes from +7 to +2 during the reaction.

(a) Both (A) and (R) are true and (R) is the correct explanation of (A).
(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).
(c) (A) is true but (R) is false.
(d) (A) is false but (R) is true.

**Ans.** (A) Here, $Cl_2$ is both oxidised and reduced in $ClO^-$ and $Cl^-$, respectively. Since $Cl^-$ cannot act as an oxidising agent (O.A.). Therefore, $Cl_2$ bleaches substances due to oxidising action of hypochlorite $ClO^-$ ion.

(B) (i) Writing the O.N. of each atom, we get: $3HCl + HNO_3 \rightarrow Cl_2 + NOCl + 2H_2O$ Oxidation number of chlorine goes from –1 (in HCl) to 0 in this case (in $Cl_2$). As a result of the oxidation of $Cl^-$, HCl is a reducing agent.

$HNO_3$ functions as an oxidising agent because the O.N. of N reduces from +5 (in $HNO_3$) to +3 (in NOCl).

Thus, reaction (i) is, therefore, a redox reaction.

(ii) $HgCl_2 + 2KI \rightarrow HgI_2 + 2KCl$

There is no change in the O.N. of any of the atoms, hence this is not a redox reaction.

(iii) $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

$Fe_2O_3$ works as an oxidising agent because the O.N. of Fe falls from +3 (in $Fe_2O_3$) to 0 (in Fe).

CO works as a reducing agent because the O.N. of C increases from +2 (in CO) to +4 (in $CO_2$). As a result, we have a redox reaction.

**(C)** (d) $Na_2SO_4$

**Explanation:** (a) Let the oxidation state of sulphur in $Na_2S$ be x.

$+2 + x = 0 \Rightarrow x = -2$

(b) Let the oxidation number of sulphur in $Na_2S_4O_6$ be x.

$\Rightarrow \quad 2 + 4x - 12 = 0$

$\Rightarrow \quad 4x = +10$

$\Rightarrow \quad x = 10/4$

$\Rightarrow \quad x = +2.5$

Thus, the oxidation number of sulphur in $Na_2S_4O_6$ is +2.5.

(c) Let the oxidation number in $Na_2SO_3$ be x.

$Na_2SO_3$

$\Rightarrow \quad 2(+1) + x + (-2) \times 3 = 0$

$\Rightarrow \quad 2 + x - 6 = 0$

$\Rightarrow \quad x = +4$

Thus, the oxidation number of sulphur in $Na_2SO_3$ is +4.

(d) Let the oxidation number of sulphur in $Na_2SO_4$ be x.

$\Rightarrow 2(+1) + x + (-2) \times 4 = 0$

$\Rightarrow \quad 2 + x - 8 = 0$

$\Rightarrow \quad x = +6$

**(D)** (c) (A) is true but (R) is false.

**Explanation:** In the reaction between potassium permanganate and potassium iodide, potassium permanganate is getting reduced where as potassium iodide is getting oxidised. Thus, potassium permanganate is acting as an oxidising agent where the oxidation state of Mn changes from +7 to +4.

$$\overset{+7}{KMnO_4} + KI \xrightarrow[\text{Base}]{\text{Weak}} \overset{+4}{MnO_2} + I_2$$

Oxidation state of Mn changes from +7 to +4.

# TOPIC 4

# TYPES OF REDOX REACTIONS

## Combination Reaction

A reaction in which two or more substances combine to form a single new substance is called a combination reaction. It is also known as a synthesis reaction.

$$A + B \rightarrow AB$$

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

## Decomposition Reaction

A reaction in which a compound breaks down into two or more simpler substances is called a decomposition reaction. All decomposition reactions are not considered as redox reactions.

$$AB \rightarrow A + B$$

$$2NaH_{(s)} \xrightarrow{\Delta} 2Na_{(s)} + H_{2(g)}$$

## Displacement Reaction

A reaction in which an ion or atom in a compound is replaced by an ion or atom of another element is called a displacement reaction. In general, it is represented by the equation.

$$X + YZ \rightarrow XZ + Y$$

The displacement reaction is divided into two categories:

(1) Metal displacement

(2) Non-metal displacement

(1) **Metal displacement:** In this type of reaction, a metal in a compound can be displaced by another metal in an uncombined state.

$$TiCl_{4(l)} + 2Mg_{(s)} \xrightarrow{\Delta} Ti_{(s)} + 2MgCl_{2(s)}$$

$$Cr_2O_{3(s)} + 2Al_{(s)} \xrightarrow{\Delta} Al_2O_{3(s)} + 2Cr_{(s)}$$

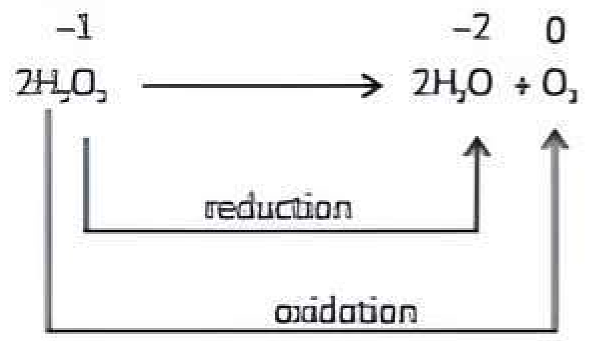
(2) **Non-metal displacement:** In this type of reaction, either a metal or non-metal will displace another non-metal of a compound present in the reaction.

$$Ca_{(s)} + 2H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$$

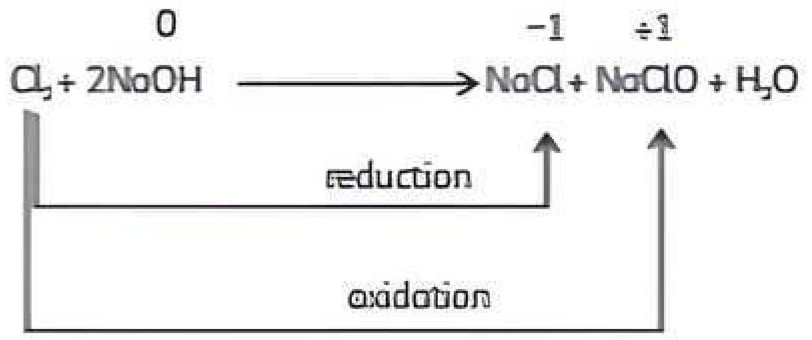
$$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$$

## Disproportionation Reaction

The reaction in which a single reactant is oxidised as well as reduced is called a Disproportionation reaction or it can also be defined as the reaction in which the oxidation state of an element is increased or decreased simultaneously. For such redox reactions, the reacting species must contain an element that has at least three oxidation states.




In the given reaction, the oxygen of hydrogen peroxide is undergoing a disproportionation reaction.




**Example 1.3:** Case Based:

*The redox reaction that involves simultaneous oxidation and reduction of atoms of the same element from one oxidation state (OS) to two different oxidation states is known as a disproportionation reaction. The disproportionation reaction is also known as the dismutation reaction. The requirement for a disproportionation reaction to occur is, the element undergoing disproportionation should exhibit a minimum of three different oxidation states and the element must be less stable in a particular oxidation state from which it can be both oxidised as well as reduced to relatively more stable oxidation states.*

(A) $MnO_4^{2-}$ undergoes a disproportionation reaction in an acidic medium but $MnO_4^-$ does not. Give reason.

(B) Disproportionation means:
(a) only oxidation
(b) only reduction
(c) both, oxidation and reduction of a reactant
(d) neither oxidation nor reduction

(C) The reaction $3ClO^-_{(aq)} \rightarrow ClO^-_{3(aq)} + 2Cl^-_{(aq)}$ is an example of:
(a) Oxidation reaction
(b) Reduction reaction
(c) Disproportionation reaction
(d) Decomposition reaction

(D) A substance which can be oxidised as well as reduced easily is:
(a) HCl (b) $Cl_2O_7$
(c) $HClO_3$ (d) $HClO_4$

(E) Assertion (A): The decomposition of hydrogen peroxide to form water and oxygen is an example of a disproportionation reaction.

Reason (R): The oxygen of peroxide is in −1 oxidation state and it is converted to zero oxidation state in $O_2$ and −2 oxidation state in $H_2O$.

(a) Both (A) and (R) are true and (R) is the correct explanation of (A).
(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).
(c) (A) is true but (R) is false.
(d) (A) is false but (R) is true.

**Ans.** (A) Disproportionation is a redox reaction in which one intermediate oxidation state component transforms into two higher and lower oxidation state compounds. Manganese's oxidation states ranging from +2 to +7 in its various compounds. $MnO_4^-$ has the maximum oxidation state of +7, hence disproportionation is impossible, but $MnO_4^{2-}$ has a +6 oxidation state, which can be oxidised as well as reduced.

(B) *(c) Both, oxidation and reduction of a reactant*

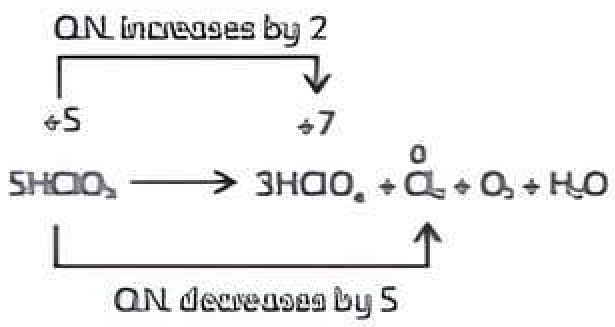
**Explanation:** In a disproportionation reaction, reactant will undergo both oxidations as well as reduction.

(C) *(c) Disproportionation reaction*

**Explanation:** The reaction $3ClO^-_{(aq)} \rightarrow ClO^-_{3(aq)} + 2Cl^-_{(aq)}$ is an example of disproportion reaction as here $ClO^-$ getting oxidised as well as reduced at the same time.

(D) *(c) $HClO_3$*

**Explanation:** $HClO_3$ will get both oxidised and reduced and thus, will show a disproportionate reaction. It can be shown from an example :




(E) *(a) Both (A) and (R) are true and (R) is the correct explanation of (A).*

**Explanation:** $2H_2O_{2(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$

Here the oxygen of peroxide, which is present in the −1 state, is converted to zero oxidation state in $O_2$ undergoing oxidation and decreases to −2 oxidation state in $H_2O$ undergoing reduction. So, it is an example of a disproportionation reaction.

## TOPIC 5

# BALANCING OF REDOX REACTION

There are two methods to balance a redox reaction

(1) Oxidation number method

(2) Half equation method or ion-electron method

## Oxidation Number Method

In this method, the number of electrons lost in oxidation must be equal to the number of electrons gained in reduction.

Following rules are followed for balancing reactions through the oxidation number method:

(1) Write the skeletal equation of all the reactants and products of the reaction.

(2) Indicate the oxidation number of each element and identify the elements undergoing changes in oxidation number.

(3) Equalise the increase or decrease in oxidation number by multiplying both reactant and products undergoing a change in oxidation number by a suitable integer.

(4) Balance all atoms other than H and O then balance the O atom by adding water molecules to the side short of atoms.

(5) In the case of ionic reactions:

**Acidic medium:** First balance O atoms by adding $H_2O$ molecules to the side deficient in atoms and then balance H-atoms by adding $H^+$ ions to the side deficient in H-atoms.

**Basic medium:** First balance O atoms by adding $H_2O$ molecules to whatever side is deficient in O atoms. The H-atoms are then balanced by adding $H_2O$ molecules equal in number to the deficiency of H-atoms and an equal number of $OH^-$ ions are added to the opposite side of the equations.

**Example 1.4: Permanganate ion reacts with bromide ion in basic medium to give manganese dioxide and bromate ion. Write a balanced Ionic equation for the reaction.** **[NCERT]**

**Ans. Step 1:** The skeletal equation is:

$$MnO_{4(aq)}^{-} + Br^{-}_{(aq)} \rightarrow MnO_{2(s)} + BrO_{3(aq)}^{-}$$

**Step 2:** Now oxidation numbers for Mn and Br

$$\overset{+7}{Mn}O_{4(aq)}^{-} + \overset{-1}{Br}{}^{-}_{(aq)} \rightarrow \overset{+4}{Mn}O_{2(s)} + \overset{+5}{Br}O_{3(aq)}^{-}$$

This indicates that the permanganate ion is the oxidant.

**Step 3:** Calculate the increase and decrease of oxidation number, and make the increase equal to the decrease.

$$2\overset{+7}{Mn}O_{4(aq)}^{-} + \overset{-1}{Br}{}^{-}_{(aq)} \rightarrow 2\overset{+4}{Mn}O_{2(s)} + \overset{+5}{Br}O_{3(aq)}^{-}$$

**Step 4:** As the reaction occurs in the basic medium, and the ionic charges are not equal on both sides, add $2OH^-$ ions on the right to make ionic charges equal.

$$2MnO_{4(aq)}^{-} + Br^{-}_{(aq)} \rightarrow 2MnO_{2(s)} + BrO_{3(aq)}^{-} + 2OH^{-}_{(aq)}$$

**Step 5:** Finally, count the hydrogen atoms and add an appropriate number of water molecules on the left side to achieve a balanced redox reaction.

$$2MnO_{4(aq)}^{-} + Br^{-}_{(aq)} + H_2O_{(l)} \rightarrow 2MnO_{2(s)} + BrO_{3(aq)}^{-} + 2OH^{-}_{(aq)}$$

### Limitation of the concept of oxidation number

According to the concept of oxidation number, oxidation means an increase in oxidation number by loss of electrons, and reduction means a decrease in oxidation number by the gain of electrons. However, during oxidation, there is a decrease in electron density while the increase in electron density around the atom undergoes reduction.

## Half Reaction Method/Ion-Electron Method

In this method, the two half equations are balanced separately and then added together to give a balanced equation. Rule for balancing the equation:

(1) Write the skeleton equation and indicate the oxidation number of all reactants and products.

(2) Find out the species which are oxidised and which are reduced.

(3) Split the skeleton equation into two half-reactions, *i.e.*, oxidation half-reaction and reduction half-reaction.

(4) Balance the two half-reaction equations separately by the rules which are described below

(i) In each half-reaction, first balance the atoms of the elements which have undergone a change in oxidation number.

(ii) Add electrons to whatever side is necessary to make up the difference in oxidation number in each half-reaction.

(5) **In the acidic medium:** H-atoms are balanced by adding $H^+$ ions to the side deficient in H-atoms. However, in the basic medium, H-atoms are balanced by adding $H_2O$ molecules equal in number to the deficiency of H-atoms and an equal number of $OH^-$ ions are included in the opposite side of the equation.

(6) The two half-reactions are then multiplied by suitable integers so that the total number of electrons gained in one half of the reaction is equal to the number of electrons lost in the other half-reaction. The two half-reactions are then added up.

(7) To verify whether the equation thus obtained is balanced or not, the total charge on either side of the equation must be equal.

**Example 1.5: Permanganate (VII) ion, in basic solution oxidises iodide ion to produce molecular iodine and manganese oxide. Write a balanced ionic equation to represent this redox reaction. [NCERT]**

**Ans.** **Step 1:** First, write the skeletal ionic equations

$$MnO_4^-{}_{(aq)} + I^-{}_{(aq)} \rightarrow MnO_{2(s)} + I_{2(s)}$$

**Step 2:** The two half-reactions are:

Oxidation half-reaction: $\overset{-1}{I^-}_{(aq)} \rightarrow \overset{0}{I}_{2(s)}$

Reduction half-reaction: $\overset{+7}{Mn}O^-_{4(aq)} \rightarrow \overset{+4}{Mn}O_{2(s)}$

**Step 3:** To balance the I atoms in the oxidation half-reaction multiply $I^-$ by 2.

$$2I^-{}_{(aq)} \rightarrow I_{2(s)}$$

**Step 4:** To balance the oxygen atoms in the reduction half-reaction, we add two water molecules on the right side:

$$MnO_4^-{}_{(aq)} \rightarrow MnO_{2(s)} + 2H_2O_{(l)}$$

To balance the H atoms, we add four $H^+$ ions on the left:

$$MnO^-_{4(aq)} + 4H^+_{(aq)} \rightarrow MnO_{2(s)} + 2H_2O_{(l)}$$

As the reaction takes place in a basic solution, therefore, we add four $OH^-$ ions to both sides of the equation:

$$MnO^-_{4(aq)} + 4H^+_{(aq)} + 4OH^-_{(aq)} \rightarrow MnO_{2(s)} + 2H_2O_{(l)} + 4OH^-_{(aq)}$$

Replace the $H^+$ and $OH^-$ ions with water and then the equation will be

$$MnO_4^-{}_{(aq)} + 2H_2O_{(l)} \rightarrow MnO_{2(s)} + 4OH^-{}_{(aq)}$$

**Step 5:** In this step, we balance the charges of the two half-reactions in the following manner

$$2I^-{}_{(aq)} \rightarrow I_{2(s)} + 2e^-$$

$$MnO_4^-{}_{(aq)} + 2H_2O_{(l)} + 3e^- \rightarrow MnO_{2(s)} + 4OH^-{}_{(aq)}$$

Now to equalise the number of electrons, multiply the oxidation half-reaction by 3 and reduction half-reaction by 2

$$6I^-{}_{(aq)} \rightarrow 3I_{2(s)} + 6e^-$$

$$2\,MnO^-_{4(aq)} + 4H_2O_{(l)} + 6e^- \rightarrow 2MnO_{2(s)} + 8OH^-_{(aq)}$$

**Step 6:** Add two half-reactions to obtain the net reactions after cancelling electrons on both sides

$$6I^-_{(aq)} + 2MnO^-_{4(aq)} + 4H_2O_{(l)} \rightarrow 3I_{2(s)} + 2MnO_{2(s)} + 8OH^-_{(aq)}$$

**Step 7:** Finally, check if the equation is balanced with respect to the number of atoms and charges on both sides.

## Redox reaction as the basis for Titration

In redox systems, the titration method can be adopted to determine the strength of a reductant/oxidant using a redox-sensitive indicator. The usage of indicators in redox titration is as follows:

(1) The reagent itself is intensely coloured, for example, permanganate ion acts as the self-indicator. The visible endpoint, in this case, is achieved after the last of the reductant ($Fe^{2+}$ or $Cr_2O_7^{2-}$) is oxidised and the first lasting tinge of pink colour appears at a concentration as low as $10^{-6}$ mol $dm^{-3}$. This ensures a minimal 'overshoot' in colour beyond the equivalence point, the point where the reductant and the oxidant are equal in terms of their mole stoichiometry.

(2) There are indicators that are oxidised immediately after the last bit of the reactant is consumed, producing a colour change. For example, $Cr_2O_7^{2-}$, which is not a self-indicator, but oxidises the indicator substance diphenylamine just after the equivalence point to produce an intense blue colour, thus signalling the end point.

(3) Its use is restricted to those reagents which are able to oxidise $I^-$ ions, say, for example Cu(II):

$$2Cu^{2+}{}_{(aq)} + 4I^-{}_{(aq)} \rightarrow Cu_2I_{2(s)} + I_{2(aq)}$$

This method relies on the fact that iodine itself gives an intense blue colour with starch and has a very specific reaction with thiosulphate ions $S_2O_3^{2-}$ which too is a redox reaction:

$$I_{2(aq)} + 2S_2O_3^{2-}{}_{(aq)} \rightarrow 2I^-{}_{(aq)} + S_4O_6^{2-}{}_{(aq)}$$

$I_2$ though insoluble in water remains in a solution containing KI as $KI_3$. In addition to starch, after the liberation of iodine from the reaction of $Cu^{2+}$ ions on iodide ions, an intense blue colour appears. This colour disappears as soon as the iodine is consumed by the thiosulphate ions. Thus, the end-point can easily be tracked and the rest is the stoichiometric calculation only.

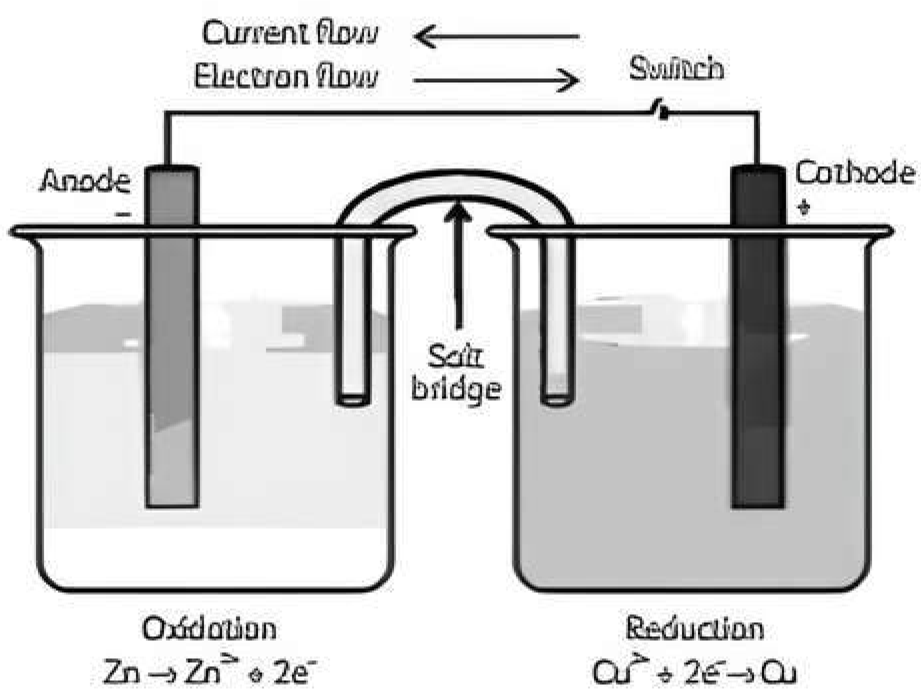
# TOPIC 6

# REDOX REACTION AND ELECTRODE PROCESSES

Redox reactions can be used in converting chemical energy into electrical energy. The cells which can be used for this process is known as an electrochemical cell. Let us discuss the role of redox reactions in these cells and some of their important terms.

A Galvanic cell or voltaic cell is a simple electrochemical cell in which a redox reaction is used to convert chemical energy into electrical energy. This means electricity can be generated with help of a redox reaction in which oxidation and reduction take place in two separate compartments. Each of them consists of a metallic conductor and is dipped in a suitable electrolytic solution of the same metal. Here, the metallic rod acts as an electrode. The solution of electrolyte is known as half-cell and half-cell has a redox couple. A redox couple is defined as a solution having reduced and oxidised form of a substance together, taking part in oxidation or reduction half-reaction. It is depicted as $M^{+n}/M$ which is oxidised form/reduced form.

To prepare a Galvanic cell two half cells are externally connected through a conducting wire and internally through a salt bridge.




Galvanic Cell

When the switch is off, then there is no reaction in either of the beaker and no current will be drawn in metallic wires. As soon as the switch is turned on:

(1) The electrons which are generated at the anode will reach the cathode.

(2) The electricity flows from one beaker to another through a salt bridge. The current flows because of the difference in potentials of two redox couples. This difference is known as cell potential or electrode potential.

Following reactions will occur:

Anodic oxidation: $Zn_{(s)} \rightarrow Zn^{+2}_{(aq)} + 2e^-$

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

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

This cell can be represented as:

$$Zn|Zn^{2+} || Cu^{2+}|Cu$$

### Important

- *(1) Electrochemical series: arrangement of $E^o$ reduction of different electrodes in increasing order of electrode potential.*
- *(2) Electrode potential: The tendency of an electrode to lose or gain electrons is called electrode potential.*
- *(3) Standard electrode potentials: a large number of electrodes have been determined using standard hydrogen electrodes as the reference electrode. By convention, the standard electrode potential ($E^o$) of the hydrogen electrode is 0.00 volts.*
- *(4) A negative ($E^o$) means that the redox couple is a stronger reducing agent than the $H^+/H_2$ couple.*
- *(5) A positive ($E^o$) means that the redox couple is a weaker reducing agent than the $H^+/H_2$ couple).*

# OBJECTIVE Type Questions

[ 1 mark ]

## Multiple Choice Questions

**1. The oxidation state of Cr in $Cr(CO)_6$ is:**

(a) 0 (b) 2

(c) 2 (d) 6

**Ans.** *(a) 0*

**Explanation:** CO (carbonyl) is a neutral ligand, hence oxidation state of Cr in $Cr(CO)_6$ is zero.

### Related Theory

- *The Ions or molecules bound to the central atom/ ion in the coordination entity are called ligands. A neutral ligand is defined as a ligand that has a neither positive nor negative charge.*

**2. The oxidation number of Mn is maximum in:**

(a) $MnO_2$ (b) $K_2MnO_4$

(c) $Mn_3O_4$ (d) $KMnO_4$

**Ans.** *(d) $KMnO_4$*

**Explanation:**

(a) $MnO_2$

$$O.N. = x + 2(-2) = 0$$
$$x = +4.$$

Mn oxidation state in $MnO_2$ is + 4.

(b) $K_2MnO_4$

$$O.N. = 2(1) + x + 4(-2) = 0$$
$$x = +6.$$

Mn oxidation state in $K_2MnO_4$ is +6.

(c) $Mn_3O_4$

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

Mn oxidation in $Mn_3O_4$ is $+\frac{8}{3}$.

(d) $KMnO_4$

$$O.N. = 1(1) + x + 4(-2) = 0$$
$$x = +7.$$

Mn oxidation state in $KMnO_4$ is +7

3. The oxidation process involves:

(a) increase in oxidation number

(b) decrease in oxidation number

(c) no change in oxidation number

(d) none of the above

**Ans.** *(a) increase in oxidation number*

**Explanation:** The oxidation process involves the addition of oxygen or an electronegative element or removal of an H/electropositive element, loss of an electrons or increase in oxidation number.

4. Consider the following reaction:

$xMnO_4^- + yC_2O_4^{2-} + zH^+ \rightarrow xMn^{2+} + 2yCO_2 + z/2H_2O$

The values of $x$, $y$ and $z$ in the reaction are, respectively:

(a) 5,2 and 16 (b) 2,5 and 8

(c) 2,5 and 16 (d) 5,2 and 8

**Ans.** *(c) 2,5 and 16*

**Explanations:** $2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$

After balancing the reaction we get

$x = 2, y = 5$ and $z = 16.$

5. Which of the following reactions does not involve in neither oxidation nor reduction?

(a) $VO^{2+} \rightarrow V_2O_3$ (b) $Na \rightarrow Na^+$

(c) $CrO_4^{2-} \rightarrow Cr_2O_7^{2-}$ (d) $Zn^{2+} \rightarrow Zn$

**Ans.** *(c) $CrO_4^{2-} \rightarrow 2Cr_2O_7^{2-}$*

**Explanations:** In $VO^{2+} \rightarrow V_2O_3$, V is reduced from +4 to +3 oxidation state.

In $Na \rightarrow Na^+$, Na is oxidised from 0 to +1 oxidation state.

In $CrO_4^{-2} \rightarrow Cr_2O_7^{-2}$, Cr remains in the same oxidation state +6.

In $Zn^{+2} \rightarrow Zn$, Zn is reduced from +2 to 0 oxidation state.

6. In which of the following compounds, an element exhibits two different oxidation states?

(a) $NH_2OH$ (b) $NH_4NO_3$

(c) $N_2H_4$ (d) $N_3H$

[NCERT Exemplar]

**Ans.** *(b) $NH_4NO_3$*

**Explanations:** Actually, $NH_4NO_3$ is an ionic compound whose ions are $NH_4^+$ and $NO_3^-$ where oxidation numbers of nitrogen in the two species are different and are shown below.

$$NH_4NO_3 \rightleftharpoons NH_4^+ + NO_3^-$$

Now, in $NH_4^+$ oxidation number of H is +1. Let $x_1$ be the oxidation number of N in $(NH_4)^+$

Then $x_1 + 1 \times 4 = 1$

From here, $x_1 = -3$

Also, in $(NO_3)^-$ oxidation number of oxygen is −2. Let $x_2$ be the oxidation number of N in $(NO_3)^-$.

Then, $x_2 - 2 \times 3 = -1$

From here, $x_2 = +5$

Hence, we find two different oxidation numbers of N in $NH_4NO_3$. They are −3 and +5.

7. Which of the following elements does not show a disproportionation tendency?

(a) Cl (b) Br

(c) F (d) I

[NCERT Exemplar]

**Ans.** *(c) F*

**Explanation:** Cl, Br and I can have oxidation states from −1 to +7. But oxidation state of F is fixed (−1) as it is the most electronegative element and does not lose electrons(except in case of HOF). Hence, it does not show a disproportionation tendency.

**Caution**

- *Except fluorine, halogens ($Cl_2$, $Br_2$, $I_2$) undergo disproportionation in an alkaline medium.*

8. Calculate the oxidation number of S in $H_2SO_4$.

(a) +7 (b) +6

(c) −1 (d) +5 [Diksha]

**Ans.** *(b) +6*

**Explanation:** In $H_2SO_4$, hydrogen exists in its usual +1 state, and oxygen exists in its −2 state. Let $x$ be the oxidation number of sulphur in sulphuric acid. Then we got

$$1 \times 2 + x + (-2)4 = 0$$
$$2 + x - 8 = 0$$
$$x - 6 = 0$$
$$x = 6$$

So, sulphur exists in the + 6 oxidation state.

9. Which of the following is not an example of a redox reaction?

(a) $CaCO_3 \rightarrow CaO + CO_2$

(b) $O_2 + 2H_2 \rightarrow 2H_2O$

(c) $Na + H_2O \rightarrow NaOH + \frac{1}{2}H_2$

(d) $MnCl_3 \rightarrow MnCl_2 + \frac{1}{2}Cl_2$

**Ans.** *(a) $CaCO_3 \rightarrow CaO + CO_2$*

**Explanation:** $\overset{+2\,+4\,-2}{CaCO_3} \rightarrow \overset{+2\,-2}{CaO} + \overset{+4\,-2}{CO_2}$

Here in this reaction, we can see that there is no change in the oxidation number of any species so it is not a redox reaction.

$$\overset{0}{O_2} + \overset{0}{2H_2} \rightarrow \overset{+1\,-2}{2H_2O}$$

Now in this reaction, oxidation number of different species is changing so reduction and oxidation both processes occur in the same reaction so it is a redox reaction.

$$\overset{0}{Na} + \overset{+1\,-2}{H_2O} \rightarrow \overset{+1\,-2\,+1}{NaOH} + \frac{1}{2}\overset{0}{H_2}$$

In this reaction, oxidation number of species changes, so it is also a redox reaction.

$$\overset{+3}{MnCl_3} \rightarrow \overset{+2\,-1}{MnCl_2} + \frac{1}{2}\overset{0}{Cl_2}$$

Here, the oxidation number of Mn changes from +3 to +2 and Cl changes from −1 to zero so it is also a redox reaction.

Hence, only option A is not a redox reaction as there is no change in the oxidation number of its species.

**10.** **Which is the correct stock notation for manganese dioxide?**

(a) Mn (I) $O_2$ (b) Mn (II) $O_2$

(c) Mn (III) $O_2$ (d) Mn (IV) $O_2$

**Ans.** *(d) Mn (IV) $O_2$*

**Explanation:** Let x be the oxidation number of Mn in $MnO_2$. Since the overall charge on the complex is zero, the sum of oxidation states of all elements in it should be equal to zero Therefore,

$$x + 2(-2) = 0$$
$$x - 4 = 0$$
$$x = +4$$

Thus, the correct notation for manganese dioxide is Mn (IV) $O_2$.

**11.** **Choose the correct option which is not involved in the process of oxidation of iron.**

(a) Formation of $Fe(CO)_5$ from Fe.

(b) At high temperature, liberation of $H_2$ from steam by iron.

(c) Rusting of iron sheets.

(d) Decolouration of blue $CuSO_4$ solution by iron.

**Ans.** *(a) Formation of $Fe(CO)_5$ from Fe.*

**Explanation:** Oxidation number of Fe in $Fe(CO)_5$ is zero. In both Fe and $Fe(CO)_5$, the oxidation state of iron is zero. Hence, no change in oxidation number. While in the other three options, there is a change in oxidation number. Thus only option (a) doesn't involve the process of oxidation of iron.

### ⚠ Caution

➥ *Students might get confused between oxidation and reduction. On the basis of oxidation number, oxidation may be defined as a reaction in which the oxidation number of an element in an atom, molecule or ion increases. Reduction may be defined as a reaction in which the oxidation number of an element in an atom, ion or molecule decreases.*

**12.** **The reduction potential of elements M, N and O are +2.46V, −1.13V, and −3.13V respectively. Which of the following order is correct regarding their reducing property?**

(a) O>N>M (b) M>O>N

(c) M>N>O (d) O>M>N

**Ans.** *(a) O>N>M*

**Explanation:** The electrode with a more negative reduction potential value is a good reducing agent and has a more reducing power.

### Related Theory

➥ *Electrode potential is the tendency of an electrode to lose or gain electrons. If the reduction takes place at the electrode, it is termed as reduction potential. A more negative value of reduction potential of an element signifies a stronger reducing agent. A positive value of reduction potential of an element signifies a weaker reducing agent.*

## Assertion-Reason (A-R)

In the following question (13-16) a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choice.

(a) Both (A) and (R) are true and (R) is the correct explanation of (A).

(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).

(c) (A) is true but (R) is false.

(d) (A) is false but (R) is true.

**13.** **Assertion (A):** Oxidation-Reduction (Redox) couple is the combination of oxidised and reduced form of a substance that is involved in Oxidation-Reduction half cell.

**Reason (R):** As in representation $E^\circ_{Fe^{3+}/Fe^{2+}}$ and $E^\circ_{Cu^{2+}/Cu^{0}}$ are two Redox couples. [Delhi Gov. QB 2022]

**Ans.** *(c) (A) is true but (R) is false.*

**Explanation:** Oxidation-Reduction (Redox) couple is the combination of oxidised and reduced form of a substance. Here, $E^\circ_{Fe^{3+}/Fe^{2+}}$ and $E^\circ_{Cu^{2+}/Cu^{0}}$, represent reduction potential values.

**14.** **Assertion (A):** $H_2SO_4$ cannot act as a reducing agent.

**Reason (R):** Sulphur cannot increase its oxidation state beyond +6.

**Ans.** *(a) Both (A) and (R) are true and (R) is the correct explanation of (A).*

**Explanation:** Maximum oxidation state of S is +6, it cannot exceed it. Therefore, it cannot be further oxidised. Thus, $H_2SO_4$ cannot act as a reducing agent.

**15.** **Assertion (A):** In a redox reaction, the oxidation number of oxidant decreases while that of reductant increases.

**Reason (R):** Oxidant gains electron(s) and reductant lose electrons.

**Ans.** (a) *Both (A) and (R) are true and (R) is the correct explanation of (A).*

**Explanation:** During a redox reaction, the oxidant gains electrons and is reduced. Hence, its oxidation number decreases also, the reductant loses electrons and is oxidised, thus its oxidation number increases.

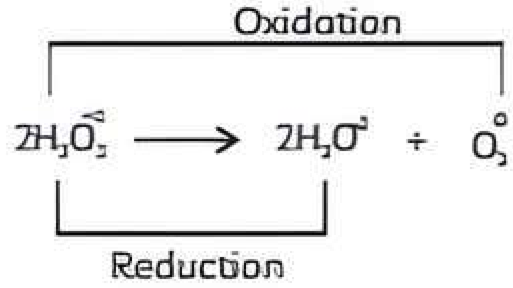
**16. Assertion (A):** The decomposition of hydrogen peroxide to form water and oxygen is an example of disproportionation reaction.

**Reason (R):** The oxygen of peroxide is in –1 oxidation state and it is converted to zero oxidation state in $O_2$ and –2 oxidation state in $H_2O$.

**[NCERT Exemplar]**

**Ans.** (a) *Both (A) and (R) are true and (R) is the correct explanation of (A).*

**Explanation:**




Therefore, the above reaction is an example of disproportionation reaction.

## CASE BASED Questions (CBQs)

[ **4** & **5** marks ]

**Read the following passages and answer the questions that follow:**

**17.** *We take copper sulphate solution in a beaker and put a copper strip or rod in it. Then on another breaker put a zinc sulphate solution and put a zinc rod or strip in it. Now reaction takes place in either of the breakers and at the interface of the metal and its salt solution in each beaker, both the reduced and oxidised forms of the same are present. A typical Galvanic cell is designed to make use of the spontaneous redox reaction between zinc and cupric ions to produce an electric current. This cell consists of a copper vessel. In which saturated $CuSO_4$ solution is filled which acts as a depolariser and dil. $H_2SO_4$ is filled which acts as an electrolyte. An amalgamated zinc rod is immersed in $ZnSO_4$. In copper vessels, there is a transparent layer all around just below the upper surface in which $CuSO_4$ crystals are kept in contact with $CuSO_4$ solution due to this the solution always remains saturated with $ZnSO_4$.*

(A) **What is a redox couple and how to represent the above experiment as a redox couple?**

(B) **Explain the observation of the above experiment when the switch is in one position. Which species will act as an oxidant?**

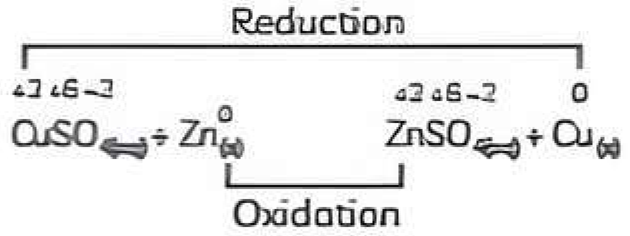
**Ans.** (A) Redox couple is defined as having together the oxidised and reduced forms of a substance taking part in an oxidation or reduction half-reaction.

$Zn^{2+} / Zn$ and $Cu^{2+} / Cu$

(B) When the switch is in the off position, no reaction takes place in either of the beakers and no current flows through the metallic wire. As soon as the switch is in the on position, we make the following observations:

(1) The transfer of electrons does not take place directly from Cu to Zn° but through the metallic wire connecting the two rods.

(2) The electricity from solution in one beaker flows by the migration of ions through the salt bridge. The flow of current is possible only when there is a potential difference between the copper and zinc known as electrodes.




Here $CuSO_4$ acts as an oxidant as it is oxidising Zn and is itself getting reduced.

**18.** *The oxidation number is the oxidation state. This results in a number of oxidation numbers denotes the oxidation state of an element in a compound. The oxidation state/metal in a compound is sometimes present according to the notation given by German chemist Alfred Stock. Which is known as stock notation. The oxidation state of an atom does not represent the "real" formal charge on that atom, or any other actual atomic property. This is particularly true of high oxidation states, where the ionisation energy required to produce a multiply positive ion is far greater than the energies available in chemical reactions. Additionally, the oxidation states of atoms in a given compound may vary depending on the choice of electronegativity scale used in* *their calculation. Thus, the oxidation state of an atom in a compound is purely a formalism. It is nevertheless important in understanding the nomenclature conventions of inorganic compounds. Also, several observations regarding chemical reactions may be explained at a basic level in terms of oxidation states.*

**(A) The oxidation state of 'S' in $KAl(SO_4)_2.12H_2O$ is:**

(a) –2 (b) –1

(c) 2 (d) +6

**(B) Name the rule in which oxidation number will be denoted in roman numbers.**

(a) Stock notation

(b) E.M.F

(c) Lewis dot rules

(d) Octet rule

**(C) Choose the right option for with respect to stock notation.**

(a) $NaCl \rightarrow 2 = Na(II)Cl$

(b) $HgCl_2 \rightarrow 1 = Hg(I)Cl_2$

(c) $AlF_3 \rightarrow 2 = Al(II)F_3$

(d) $SiCl_4 \rightarrow 3 = Si(III)Cl_4$

**(D) A reagent which can increase the oxidation number of an element in the given substance.**

(a) Oxidation

(b) Reducing reagent

(c) Oxidising reagent

(d) Reduction

**(E) Which of the following statements regarding $HClO_4$ and $HClO_3$ is true?**

(a) The O.N. for chlorine in $HClO_4$ has increased in $HClO_3$

(b) The O.N. of oxygen in $HClO_4$ has been decreased in $HClO_3$

(c) The O.N. of chlorine in $HClO_4$ has been decreased in $HClO_3$

(d) The oxidation numbers for all atoms are the same in both molecules

**Ans.** (A) *(d) +6*

**Explanation:** Let the oxidation number of Sulphur be $x$ in $KAl(SO_4)_2.12H_2O$. Oxidation no. of known elements are:

$K = +1, Al = +3, O = -2, H = +1$

$+1 + 3 + 2[x + 4(-2)] = 0$

$+4 + 2x - 16 = 0$

$2x = 12$

$x = +6$

Hence, the oxidation no. of S is +6.

(B) *(a) Stock notation*

**Explanation:** Stock notation is the oxidation number expressed by putting a roman numeral representing the oxidation number in parenthesis after the symbol of the metal in the molecular formula.

(C) *(b) $HgCl_2 \rightarrow 1 = Hg(I)Cl_2$*

**Explanation:** The oxidation number of Hg in $HgCl_2$ is 1. Thus according to stock notation, it can be represented as

$$HgCl_2 \rightarrow 1 = Hg(I)Cl_2$$

(D) *(c) Oxidising reagent*

**Explanation:** A reagent which can increase the oxidation number of an element in the given substance is known as oxidising reagent

(E) *(c) The O.N. of chlorine in $HClO_4$ has been decreased in $HClO_3$*

**Explanation:**

(1) Atoms in elemental form have an oxidation state of 0.

(2) Halogens are commonly given an oxidation state of –1.

(3) Hydrogen and alkali metals are commonly given an oxidation state of +1.

(4) Oxygen is commonly given an oxidation number of –2.

We can assume that the hydrogens and oxygens in the two compounds preserve the same oxidation state by keeping these criteria in mind. Yet, in $HClO_3$, chlorine is bound to one less oxygen. Comparing $HClO_4$ to $HClO_3$, the oxidation state of Cl has therefore fallen from +7 to +5.

**19.** *Oxidation was used to describe the addition of oxygen to an element or a compound. Because of the presence of dioxygen in the atmosphere (~20%). Many elements combine with it and this is the principal reason why they commonly occur on the earth in the form of their oxides. Oxidation is the loss of electrons during a reaction by a molecule, atom or ion. Oxidation occurs when the oxidation state of a molecule, atom or ion is increased. The opposite process is called reduction, which occurs when there is a gain of electrons or the oxidation state of an atom, molecule, or ion decreases.*

**(A) The oxidation process involves:**

(a) Decrease in oxidation number

(b) Increase in oxidation number

(c) No change in oxidation number

(d) None of the above

**(B) The addition of oxygen or electronegative element to a substance:**

(a) Reduction

(b) Oxidising agent

(c) Reducing agent

(d) Oxidation.

**(C) Select the correct option for oxidation example.**

(a) $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

(b) $Zn^{2+}/Zn$ and $Cu^{2+}/Cu$.

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

(d) $Cr_2O_{3(s)} + 2Al_{(s)} \xrightarrow{\Delta} Al_2O_{3(s)} + 2Cr_{(s)}$

(D) Which of the following may act both as an oxidising and a reducing agent?

(a) $H_2O_2$ (b) $MnO_2$

(c) $SO_2$ (d) All of these

**Ans.** (A) *(b) Increase in oxidation number*

**Explanation:** Oxidation is that process which results in an increase of the oxidation number.

(B) *(d) oxidation*

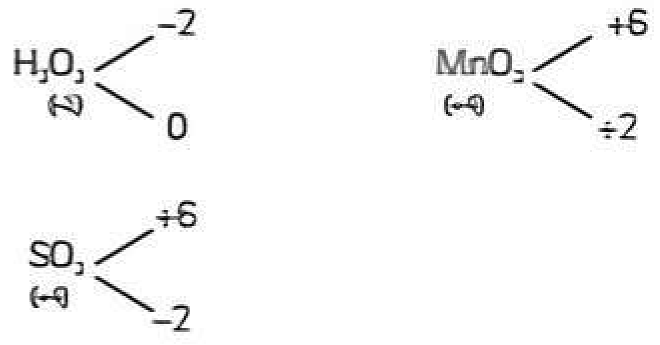
**Explanation:** It is defined as the addition of oxygen or an electronegative element to a substance or the removal of hydrogen/ electropositive element from a substance.

(C) *(a)* $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

**Explanation:** In this reaction, the addition of oxygen to magnesium took place which represents oxidation.

(D) *(d) All of these*

**Explanation:** All the given species can act as both oxidising and reducing agents.




**Related Theory**

- *If an element is in its intermediate oxidation state in a compound, the compound can function both as an oxidising agent as well as a reducing agent as the element can increase or decrease its oxidation state.*

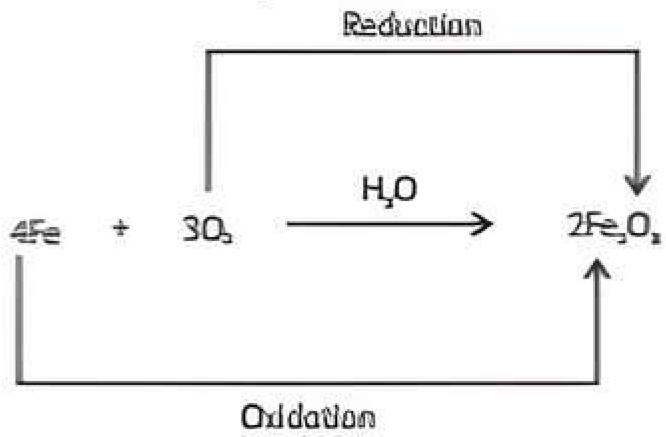
**20.** *Place a strip of metallic zinc in an aqueous solution of copper nitrate for about one hour. We can observe that the strip becomes coated with reddish metallic copper and the blue colour of the solution disappears. Formation of* $Zn^{2+}$ *ions among the product can easily judge when the blue colour solution due to the* $Cu^{2+}$ *has disappeared. Redox reactions are characterised by the actual or formal transfer of electrons between chemical processes, most often with one species undergoing oxidation (losing electrons) while another species undergoes reduction. The chemical species from which the electron is removed is said to have been oxidised, while the chemical species to which the electron is added is said to have been reduced.*

**(A) Explain the redox reaction with the help of an example.**

**(B) Explain the experiment with the help of an equation.**

**(C) Why does the metallic zinc strip turn reddish in colour?**

**Ans.** (A) Redox reaction is defined as the reaction in which oxidation and reduction take place simultaneously.



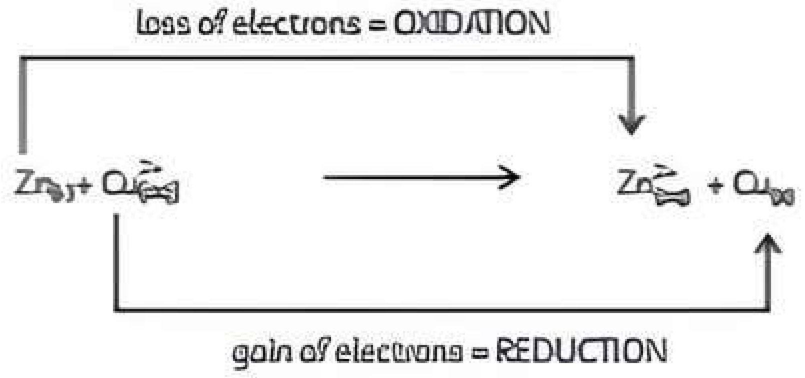

(B) Oxidation half-reaction:

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

Reduction half-reaction:

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

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




(C) This is because copper ions from copper nitrate solution get reduced by accepting electrons from zinc. This causes deposition of copper on zinc metal strip which is red in colour.

## VERY SHORT ANSWER Type Questions (VSA)

[ 1 mark ]

**21.** The reaction:

$Cl_{2(g)} + 2OH^-_{(aq)} \rightarrow ClO^-_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$ represents the process of bleaching. Identify and name the species that bleaches the substances due to their oxidising action. [NCERT Exemplar]

**Ans.** Hypochlorite ion.

**Explanation:**

$$\overset{0}{Cl}_{2(g)} + 2\overset{-2+1}{OH}_{(aq)} \rightarrow \overset{+1-2}{ClO}_{(aq)} + \overset{+1-2}{H_2O}_{(l)} + \overset{-1}{Cl}$$

In this reaction, the O.N. of Cl increases from 0 to 1 as well as decreases from 0 to −1. So it acts as both a reducing as well as an oxidising agent. This is an example of a disproportion reaction. in this reaction $ClO^-$ species bleaches the substance due to its oxidising action. Hypochlorite ions can decrease their oxidation number from +1 to 0 or −1.

**22. The standard reduction electrode potential of three metals: X, Y and Z are −1.2V, + 0.5V and −3.0V respectively. What will be the order of reducing power?**

**Ans.** As we know that the higher the reduction potential, the lesser is reducing power and *vice versa*.

Therefore, the reducing power will be Z>X>Y

**23. Explain the application of redox reaction.**

**Ans.** The redox reaction is used for electroplating by applying a thin coat of a material on an object. in the production of gold-plated jewellery.

**24. Calculate the oxidation number of S in $H_2S$.**

**Ans.** Sum of the oxidation of atoms in

$$H_2S = 2(+1) + x$$
$$= 2 + x$$

But the sum of the oxidation numbers of various atoms in $H_2S$ is zero.

Therefore, $2 + x = 0$ or $x = -2$

Thus, oxidation number of S in $H_2S$ is −2.

## SHORT ANSWER Type-I Questions (SA-I)

[ 2 marks ]

**25. Nitric acid is an oxidising agent and reacts with PbO but it does not react with $PbO_2$. Explain why? [NCERT Exemplar]**

**Ans.** PbO is a basic oxide and simple acid basic reaction takes place between PbO and $HNO_3$. On the other hand, in $PbO_2$ lead is in +4 oxidation state and cannot be oxidised further. Therefore, no reaction takes place.

$2PbO + 4HNO_3 \rightarrow 2Pb(NO_3)_2 + 2H_2O$ (Acid base reaction)

**26. Write formula for the following compounds:**

**(A) Mercury (II) chloride**

**(B) Nickel (II) sulphate**

**(C) Iron (III) sulphate**

**(D) Chromium (III) oxide**

**Ans.** (A) $HgCl_2$

(B) $NiSO_4$

(C) $Fe_2(SO_4)_3$

(D) $Cr_2O_3$

**27. What is the fractional oxidation number?**

**Ans.** When the same element is involved in different bonding in a species, their actual oxidation states are whole numbers but an average of these is fractional. This oxidation is known as Fractional Oxidation Number. Elements as such do not have any fractional oxidation numbers. But in some cases, they exist.

**28. Define standard electrode potential.**

**Ans.** Standard electrode potential is a measurement of the potential for equilibrium. There is a potential difference between the electrode and the electrolyte called the potential of the electrode. When unity is the concentration of all the species involved in a half-cell, the electrode potential is known as the standard electrode potential.

**Related Theory**

→ *Under standard conditions, the standard electrode potential occurs in an electrochemical cell, say the temperature = 298K, pressure = 1atm, concentration = 1M. The symbol '$E^\circ$ cell' represents the standard electrode potential of a cell. By convention, the standard electrode potential ($E^\circ$) of the hydrogen electrode is 0 volts.*

**29. PbO and $PbO_2$ react with HCl according to the following chemical equations:**

**(i) $2PbO + 4HCl \rightarrow 2PbCl_2 + 2H_2O$**

**(ii) $PbO_2 + 4HCl \rightarrow 2PbCl_2 + Cl_2 + 2H_2O$**

**Why do these compounds differ in their reactivity? [NCERT Exemplar]**

**Ans.** In reaction (i), none of the atom's oxidation numbers changes. Therefore, it is not a redox reaction. It is an acid-base reaction because PbO is a basic oxide which reacts with HCl acid. The reaction (ii) is a redox reaction in which $PbO_2$ gets reduced and acts as an oxidising agent.

**30. The following two reactions can occur during electrolysis of aqueous sodium chloride solution.**

$$Na^+ + e^- \rightarrow Na, E^\circ = -2.71V$$

$$2H_2O + 2e^- \rightarrow H_2 + 2OH^-, E^\circ = -0.83V$$

**Which reaction takes place preferentially and why?**

**Ans.** Since $E^\circ$ of $H_2O$ is less negative than that of $Na^+$ ions, therefore, reduction of $H_2O$ takes place preferentially to evolve $H_2$ gas at the cathode.

## SHORT ANSWER Type-II Questions (SA-II)

[ 3 marks ]

**31. $MnO_4^{2-}$ undergoes disproportionation reaction in acidic medium but $MnO_4^-$ does not. Give reason. [NCERT Exemplar]**

**Ans.** In $MnO_4^-$, Mn is in the highest oxidation state i.e. +7. Therefore, it does not undergo disproportionation. $MnO_4$ has oxidation number +6 which can both be increased or decreased, thus it undergoes disproportionation as follows:

$$3\overset{+6}{Mn}O_4^{2-} + 4H^+ \rightarrow 2\overset{+7}{Mn}O_4^- + \overset{+4}{Mn}O_2 + 2H_2O$$

**32. Explain with suitable reasons:**

**(A) Reaction $FeSO_{4(aq)} + Cu_{(s)} \rightarrow CuSO_{4(aq)}$ + Fe does not occur.**

**(B) Zinc can displace copper from aqueous $CuSO_4$ solution but Ag cannot.**

**(C) Solution of $AgNO_3$ turns blue when copper rod is immersed in it.**

**[Delhi Gov. QB 2022]**

**Ans.** (A) The standard reduction values for $Fe^{2+}$ and $Cu^{2+}$ are given as:

$$E^\circ_{Fe^{2+}/Fe} = -0.44\ V$$

$$E^\circ_{Cu^{2+}/Cu} = +0.34\ V$$

∴ As we know that a metal having a more negative value of E° can displace the metal having a less negative or positive value of E° from their salt solution. So, Cu cannot displace Fe from $FeSO_4$ solution.

(B) The metal zinc (Zn) displaces copper (Cu) from copper sulphate $CuSO_4$ because zinc is more reactive than copper. In contrast, silver is not able to displace copper from copper sulphate $CuSO_4$ because it is less reactive than copper.

(C) As we know that a metal having a more negative value of E° can displace the metal having a less negative or positive value of E° from its salt solution. So let's check their reduction potentials:

$$E^\circ_{Ag^+/Ag} = 0.799\ V$$

$$E^\circ_{Cu^{2+}/Cu} = +0.34\ V$$

Thus, Cu will displace Ag from $AgNO_3$ solution and the colour of solution will turn blue.

$$Cu^{+2} + AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$$

**33. Calculate the oxidation number of phosphorous in the following species.**

**(A) $HPO_3^{2-}$ (B) $PO_4^{3-}$**

**[NCERT Exemplar]**

**Ans.** Let the oxidation number of phosphorous be $x$

Oxidation number of O atom $= -2$

Oxidation number of H atom $= +1$

(A) $HPO_3^{2-}$:

$$+1 + x + (-2) \times 3 = -2$$
$$+1 + x - 6 = -2$$
$$x - 5 = -2$$
$$x = -2 + 5$$
$$x = +3$$

Thus, oxidation state of phosphorous is +3.

(B) $PO_4^{3-}$:

$$x + (-2) \times 4 = -3$$
$$x - 8 = -3$$
$$x = -3 + 8$$
$$x = +5$$

Thus, oxidation state of phosphorous in this case is +5.

**34. Explain redox reactions on the basis of electron transfer. Give a suitable example. [NCERT Exemplar]**

**Ans.** When a redox reaction takes place, the species which loses the electrons is undergoing an oxidation reaction. It, therefore, acts as an oxidising agent or oxidant. The species that accepts electrons is undergoing a reduction reaction and therefore it behaves as a reductant. The above-mentioned transfer is largely based on the relative electronegativity difference between two interacting species. The higher electronegative element accepts electrons while on the other hand, the electropositive element loses electrons. For example:

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

Here, from the above reaction, it can be seen that sodium is oxidised as it loosing two electrons while hydrogen is reduced as it accept two electrons. Therefore, it is a redox reaction.

# LONG ANSWER Type Questions (LA)

[ 4 & 5 marks ]

**35. Explain different types of redox reactions with examples.**

**Ans.** Different types of reaction:

**Combination reactions:** A reaction in which two or more substances combine to form a single new substance. It is also known as synthesis reactions.

$$A + B \rightarrow AB$$

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2\,NaCl_{(s)}$$

**Decomposition reactions:** A reaction in which a compound breaks down into two or more simpler substances.

$$AB \rightarrow A + B$$

$$2NaH_{(s)} \xrightarrow{\Delta} 2Na_{(s)} + H_{2(g)}$$

**Displacement reactions:** A reaction in which an ion in a compound is replaced by an ion of another element.

$$X + YZ \rightarrow XZ + Y$$

The displacement reaction is divided into two categories:

(1) Metal displacement

(2) Non-metal displacement

(1) **Metal displacement:** A metal in a compound can be displaced by another metal in an uncombined state.

$$TiCl_{4(l)} + 2Mg_{(s)} \xrightarrow{\Delta} Ti_{(s)} + 2MgCl_{2(s)}$$

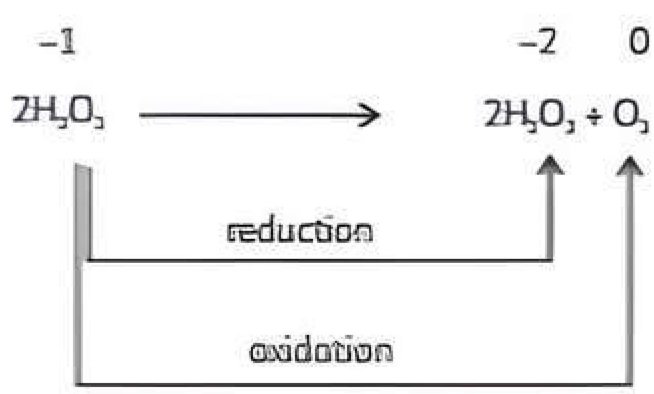
$$Cr_2O_{3(s)} + 2Al_{(s)} \xrightarrow{\Delta} Al_2O_{3(s)} + 2Cr_{(s)}$$

(2) **Non-metal displacement:** In this type of reaction, either a metal or non-metal will displace another non-metal of a compound present in the reaction.

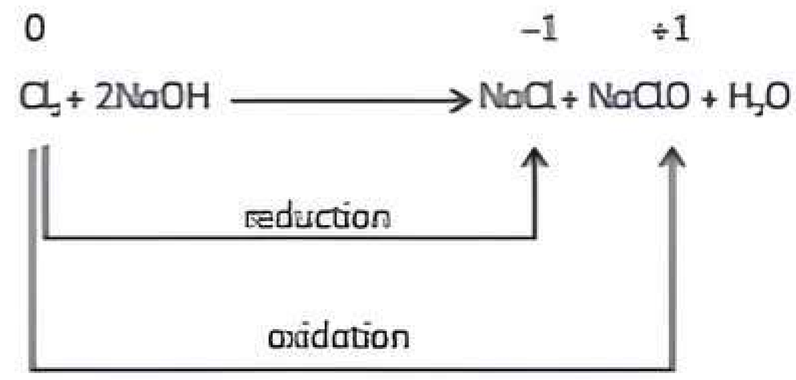
$$Ca_{(s)} + 2H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$$

**Disproportionation reactions:** The reaction in which a single reactant is oxidised and reduced.

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




$$Cl_2 + 2NaOH \rightarrow NaCl + NaClO + H_2O$$




**36. Balance the following equations by oxidation number method:**

$$CuO + NH_3 \rightarrow Cu + N_2 + H_2O$$

$$K_2MnO_4 + H_2O \rightarrow MnO_2 + KMnO_4 + KOH$$

**Ans.** $CuO + NH_3 \rightarrow Cu + N_2 + H_2O$

Let us observe the chemical equation

$$\overset{+2}{Cu}O + \overset{-3}{N}H_3 \rightarrow \overset{0}{Cu} + \overset{0}{N_2} + H_2O$$

The oxidation number of copper decreases from +2 to 0 and that of nitrogen increases from −3 to 0.

To balance, there should be three atoms of copper and two atoms of nitrogen

$$3CuO + 2NH_3 \rightarrow 3Cu + N_2 + H_2O$$

To balance hydrogen and oxygen

$$3CuO + 2NH_3 \rightarrow 3Cu + N_2 + 3H_2O$$

$$K_2MnO_4 + H_2O \rightarrow MnO_2 + KMnO_4 + KOH$$

Let us observe the chemical equation:

$$K_2\overset{+6}{Mn}O_4 + H_2O \rightarrow \overset{+4}{Mn}O_2 + K\overset{+7}{Mn}O_4 + KOH$$

The oxidation number of manganese changes from +6 to +4, in one mole, and in the other mole, the oxidation number changes from +6 to +7. 1 mol acquires two electrons while the other losses 1 electron. To balance the oxidation number of manganese, multiply it by 2.

$$2K_2MnO_4 + H_2O \rightarrow MnO_2 + 2KMnO_4 + KOH$$

Further balancing the equation, we have

$$3K_2MnO_4 + 2H_2O \rightarrow MnO_2 + 2KMnO_4 + 4KOH$$

**37. Explain Galvanic cell.**

**Ans.** (1) A Galvanic cell or voltaic cell is a simple electrochemical cell in which a redox reaction is used to convert chemical energy into electrical energy.

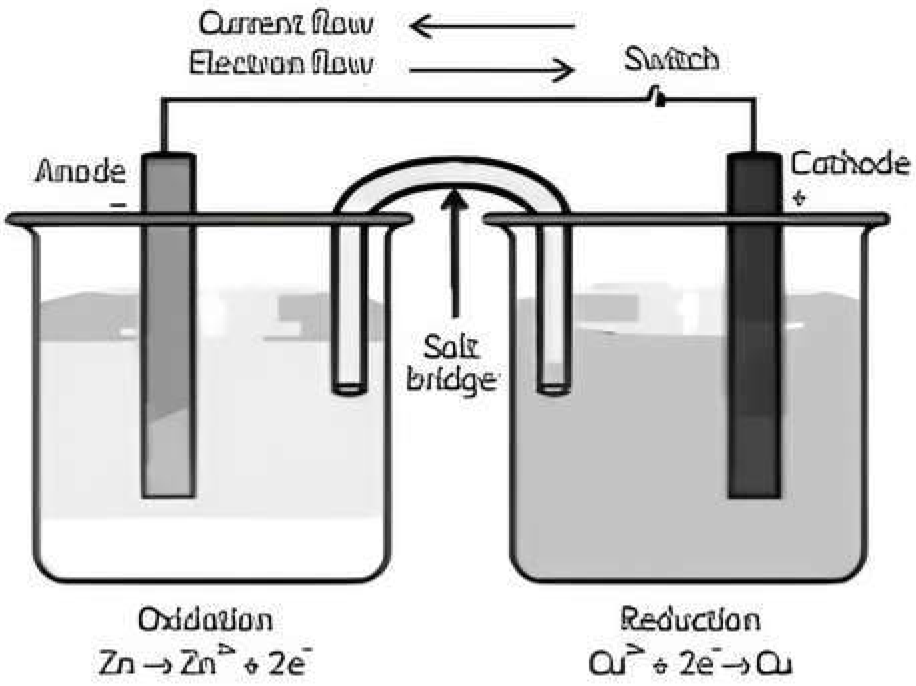
(2) This means electricity can be generated with help of a redox reaction.

(3) In which oxidation and reduction take place in two separate compartments. Each of them consists of a metallic conductor and is dipped in a suitable electrolytic solution of the same metal. Here, the metallic rod acts as an electrode. The solution of electrolyte is known as half-cell and half-cell has a redox couple.

(4) A redox couple is defined as a solution having reduced and oxidised forms of a substance together, taking part in oxidation or reduction half-reaction.

(5) It is depicted as $M^{+n}/M$ is oxidised form/ reduced form.

(6) To prepare a Galvanic cell two half cells are externally connected through a conducting wire and internally through a salt bridge.




*Galvanic Cell*

Anodic oxidation: $Zn \rightarrow Zn^{+2}_{(aq)} + 2e^-_{(s)}$

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

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

This cell can be represented as

$Zn|Zn^{2+} || Cu^{2+}|Cu$

## NUMERICAL Type Questions

**38. Calculate the oxidation number of:**

**(A) C in $CH_3COOH$**

**(B) All atoms in $KClO_4$** **(2m)**

**Ans.** (A) Oxidation number of oxygen = − 2

Oxidation number of oxygen = + 1

Let oxidation number of carbon = $x$

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

(B) The oxidation number of K = + 1

The oxidation number of Cl = $x$

The oxidation number of O = −2

$$1 + x - 8 = 0$$
$$x - 7 = 0$$
$$x = 7$$

(K = +1, O = − 2, Cl = 7)

**39. Calculate the oxidation number of Sb atom in $Sb_2O_5$.** **(2m)**

**Ans.** $Sb_2O_5$: Oxygen in common oxide has an oxidation state of −2.

Therefore, if $x$ is the oxidation number of Sb in $Sb_2O_5$ then,

$$2 \times x + 5 \times (-2) = 0$$
$$2x = 10$$
$$x = \frac{10}{2}$$
$$x = 5$$

The oxidation number of Sb in $Sb_2O_5$ = + 5

**40. Predict whether zinc reacts with 1M $H_2SO_4$ to give out hydrogen or not. Given that the standard potential of zinc is −0.76V.** **(2m)**

**Ans.** To predict the reaction between zinc with $H_2SO_4$. If zinc reacts, then the following reaction will take place:

$$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$$

i.e. $$Zn + 2H^+ \rightarrow Zn^{2+} + H_2$$

By conversion, the cell is represented as $Zn|Zn^{2+} || H^+ | H_2$

Standard EMF of the cell

$$E^o_{cell} = E^o_{H^+/H_2} - E^o$$
$$= 0 - (-0.76) = +0.76V$$

Since the EMF come out to be positive. Hence the reaction takes place.