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CBSE

CLASS-11th

THE CENTRAL BOARD OF SECONDARY EDUCATION

CHEMISTRY-II

SIERRA
INNOVATIONS

CONTENT

CHAPTER-8
REDOX REACTION

Content

- ✓ Concepts of Oxidation and Reduction.
- ✓ Oxidising and Reducing agent.
- ✓ Oxidation Number: Definition and Rules governing oxidation number.
- ✓ Redox Reactions and its Types.
- ✓ Methods for Balancing Redox Reactions.
- ✓ Electrochemical series and Redox Reaction as the basis for titrations.
- ✓ Important Questions

Oxidation and Reduction

Oxidation	Reduction
1.) Addition of oxygen e.g. $2\ Mg + O_2 \rightarrow 2\ MgO$	1.) Removal of oxygen e.g. Cuo +C → Cu +CO
2.) Removal of Hydrogen e.g. $H_2S + Cl_2 \rightarrow 2HCl + S$	2.) Addition of Hydrogen e.g. $S + H_2 \rightarrow H_2S$
3.) Increase in positive charge e.g. $Fe^{2+} \rightarrow Fe^{3+}$ **+ e-**	3.) Decrease in positive charge e.g. $Fe^{3+} + e^- \rightarrow Fe^{2+}$
4.) Removal of Electron e.g. $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$	4.) Addition of Electron e.g. $Fe^{3+} + e^- \rightarrow Fe^{2+}$

Note: The no. of electrons lost in oxidation process = the no. of electrons gained during reduction process.

Oxidizing and Reducing agent

Oxidizing agent

The substance which undergoes reduction itself and oxidizes other is called Oxidizing agent (Also known as **oxidant**). For eg: $K_2Cr_2O_7$, $KMnO_4$, H_2O_2, Cl_2.

Reducing agent

The substance which undergoes oxidation itself and reduces other is called Reducing agent (Also known as **reductant**). For eg: H_2, H_2S, Mg, $SnCl_2$.

Oxidation Number

- The total number of electrons that an atom either gains or losses in order to form a chemical bond with another atom.

or

- The formal charge present on an atom in a particular compound determined by certain arbitrary rule.

Rules governing oxidation number

- Fluorine is the most electronegative atom. It always has oxidation number equal to -1 in all compounds.
- In general, and as well as in oxides, oxygen atom has oxidation number equal to -2.

In case of

- Peroxide (e.g. H_2O_2, Na_2O_2) is -1
- superoxide (e.g. KO_2) is -1/2
- ozonide (e.g. KO_3) is -1/3
- In OF_2 is +2& in O_2F_2 is +1
- In general, H atom has oxidation number equal to +1. But in Metallic hydrides (e.g. NaH, KH), it is -1.
- In general, all Halogen atoms (Cl, Br, I) have oxidation number equal to -1, But if Halogen atoms is attached with a more electronegative atom that halogen, then it will show positive oxidation numbers.

Metals

- Alkali metal (Li, Na, k…) always have oxidation number +1.
- Alkaline earth metal (Be, Mg, Ca…) always have oxidation number+2.
- Aluminium always has +3 oxidation numbers.

Note:

i) Metal may have negative or zero oxidation number.
ii) Oxidation number of an element in Free State is always zero.
iii) Sum of the oxidation number of atoms of all elements in an ion is equal to the charge to the ion.
iv) Sum of the oxidation number of atoms of all elements in a molecule is zero. Oxidation number and Acid Strength

The greater the O.N. of the Element in oxyacids, the greater is Acid Strength.

$$\overset{+7}{HClO_4} > \overset{+5}{HClO_3} > \overset{+3}{HClO_2} > \overset{+1}{HClO}$$

Calculation/ determination of oxidation number of underlined element in some compounds

- $K_2Cr_2O_7$

  Let the O.N. of Cr be x then

  $2 \times (+1) + 2 \times (x) + 7 \times (-2) = 0$

  $2 + 2x - 14 = 0$ $x = +6$

- $KMnO_4$

  Let the O.N. of Mn be x the

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

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

Redox reactions

A redox (or oxidation-reduction) reaction is a type of chemical reaction that involves a transfer of electrons between two species. We can tell there has been a transfer of electrons if there is any change in the oxidation number between the reactants and the products.

or

These are those reactions comprising of simultaneous oxidation and reduction and called oxidation - Reduction or Redox reactions.

$SnCl_2 + 2HgCl_2 \longrightarrow SnCl_4 + Hg_2Cl_2$

Types of redox reactions

- Decomposition Reaction
- Combination Reaction
- Displacement Reaction
- Disproportionation Reactions

Decomposition Reaction

This kind of reaction involves the breakdown of a compound into different compounds. Examples of these types of reactions are:

- $2NaH \rightarrow 2Na + H_2$
- $2H_2O \rightarrow 2H_2 + O_2$
- $Na_2CO_3 \rightarrow Na_2O + CO_2$

All the above reactions result in the breakdown of smaller chemical compounds in the form of AB → A + B

But, there is a special case that confirms that all the **decomposition reactions** are not redox reactions. For example $CaCO_3 \rightarrow CaO + CO_2$

Combination Reaction

These reactions are the opposite of decomposition reaction and hence involve the combination of two compounds to form a single compound in the form of

A + B → AB. For example:

- $H_2 + Cl_2 \rightarrow 2HCl$
- $4Fe + 3O_2 \rightarrow 2Fe_2O_3$

Displacement Reaction

In this kind of reaction, an atom or an ion in a compound is replaced by an atom or an ion of another element. It can be represented in the form of

X + YZ → XZ + Y.

Further **displacement reaction** can be categorized into

- Metal displacement Reaction
- Non-metal displacement Reaction

Metal Displacement

In this type of reaction, a metal present in the compound is displaced by another metal. These types of reactions find their application in metallurgical processes where pure metals are obtained from their ores.

For example $CuSO_4 + Zn \rightarrow Cu + ZnSO_4$

Non-Metal Displacement

In this type of reaction, we can find a hydrogen displacement and sometimes rarely occurring reactions involving oxygen displacement.

Disproportionation Reactions

The reactions in which single reactant is oxidized and reduced is known as Disproportionation reactions.

For eg: $P_4 + 3NaOH + 3H_2O \rightarrow 3NaH_2PO_2 + PH_3$

- **Intermolecular redox reactions**

In this case one substance is oxidised and another is reduced.

$4\,HCl + MnO_2 \longrightarrow MnCl_2 + Cl_2 + 2H_2O$

Here HCl is oxidised and MnO_2 is reduced.

Disproportion - In this case the same substance is oxidised and reduced e.g.

$4KClO_3\ (+5) \rightarrow 3KClO_4\ (+7) + KCl\ (-1)$

- **Intermolecular redox reactions**

In this case one element of the compound is reduced while another element of the same compound is oxidised

$(NH_4)_2CrO_7 \rightarrow N_2 + CrO_3 + 4H_2O$

Cr is reduced and N is oxidized.

Methods for Balancing Redox Reaction

For Balancing a Chemical Equation, the two important methods are:

Oxidation Method

The certain Rules are as follows:

i) Assign oxidation number to the atoms showing a change in oxidation state.
ii) Balance the total number of atoms undergoing change in oxidation state.
iii) Balance the number of electrons gained or lost.
iv) Balance [O] on both sides by adding H_2O.

v) Balance H atoms by adding H^+ions.
vi) If the reaction proceeds in basic solution add sufficient number of OH^- ions on both sides.

Ion Electron Method

The Rules are as follows:

i) Split up the reaction into two half reactions showing oxidation and reduction separately.
ii) Balance number of atoms undergoing the change of oxidation state.
iii) Balance O on both sides by adding H_2O.
iv) Balance H atoms by adding H^+ ions.
v) Balance charge by adding required number of electrons.
vi) Make the number of electrons equal in two half reactions by multiplying with suitable coefficient.
vii) Add the two half reactions.

Example: Let us consider the skeletal equation:

$Fe^{2+} + Cr_2O7^{2-} \rightarrow Fe^{3+} + Cr^{3+}$

Step 1: Separate the equation in to two halves:

Oxidation half reaction: $Fe^{2-} \rightarrow Fe^{3+}$

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

Step 2: Balance the atoms other than hydrogen and oxygen in each half reaction individually. Here the oxidation half reaction is already balanced with respect to Fe atoms .For the reduction half reaction; we multiply the Cr^{3+} by 2 to balance Cr atoms.

Step 3: For reactions occurring in acidic medium, add water molecules to balance oxygen atoms and hydrogen ions are balanced by adding H atoms. Thus, we get:

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

Step 4: Add electrons to one side of the half reaction to balance the charges .if needed make the number of electrons equal in two half reactions by multiplying one or both half reaction by suitable coefficient.

The oxidation half reaction is thus written again to balance the charge .Now in the reduction half reaction there are 12 positive charges on the left hand side and only 6 positive charge on right hand side .Therefore, we add six electrons to left hand side .

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

To equalize the number of electrons in both reactions, we multiply oxidation half reaction by 6 and write as:

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

Step 5: We add the two half reactions to achieve the overall reaction and cancel the electrons on each side .This give us net ionic equation:

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

Step6: Verify that the equation contains the same type and number of atoms and the same charges on both sides of the equation. This last check reveals that the equation is fully balanced with respect to number atoms and the charge.

Electrochemical Series and Redox Reaction as the basis for Titrations

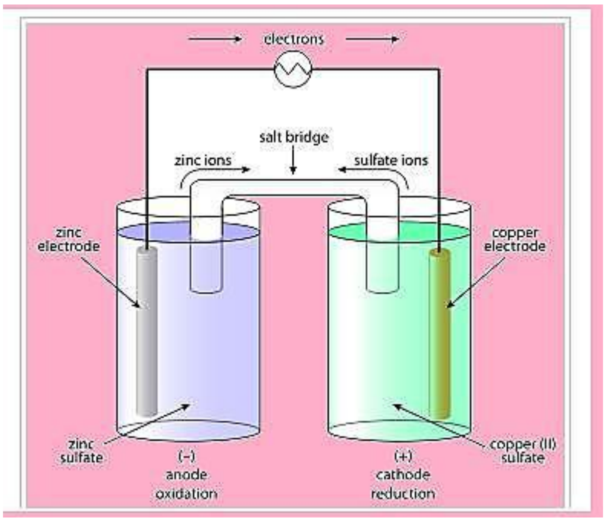
1. **Electrochemical series**

- It is the arrangement of elements in order of increasing potential. It is the series has the values starting from Negative to positive.
- Electrochemical cell is the cell in which chemical energy gets converted to electric energy.
- In it indirect redox reactions takes place.
- These reactions are spontaneous that is free energy change for this reaction is negative.

This cell consists of two half cells.

In one half cells, there is an aqueous 1molar Zinc sulphate solution with Zinc rod dipped in it.

In other half cell, there is a 1 molar aqueous solution of Copper sulphate solution with Copper rod dipped in it.



These electrodes by means of wire are attached to galvanometer.

A U-shaped tube is taken, which is sealed from both the ends with cotton plug.

In this, the electrolyte that is inert electrolyte is taken like Potassium nitrate, Ammonium nitrate etc. The electrolyte present is in semi-liquid state.

Observations -

With time we see that Zinc rod loses weight, as it has more tendency to loose electrons that is:

Zn -2 electrons → Zn^{2+} (Oxidation)

Zinc Zinc Ion

These electrons released by zinc, travel to another beaker by means of wire. In doing so, they cause deflection in galvanometer and produce current. This current travel in the direction opposite to the flow of electrons.

These electrons move to another half cell, where copper ions gain these electrons that is reduction occur. As a result, copper metal start depositing on electrode. The reaction that occurs is shown below:

Cu^{2+} + 2electrons → Cu(reduction)

Copper ions Copper Metal

The overall reaction that takes place is:

Zn + Cu^{2+} --> Zn^{2+} + Cu

Zinc Copper Zinc Ion Copper Metal

Representation of the cell

Formula to calculate standard electrode potential of cell (when concentration of electrolyte is 1 molar)

$E = E_c - E_a$ (where 'c' is cathode and 'a' is anode).

2. **Redox Reaction as the basis for Titrations**

Titration is the process in which the solutions of two reagents are allowed to react with each other.

Procedure:

- In it, one solution (known volume) is taken in Burette and the solution is called titrant.
- The other reagent is taken in flask called titration flask and the solution is called as analyte.
- The titration is carried out till both the reagents mix completely.
- The stage at which both the reagent mix completely is called endpoint.
- The endpoint is detected by an indicator.
- The objective of these titrations is to find out the exact amount of an acid (or the base) present in a given solution by reacting it against the solution of a standard base (or an acid).

Electrode potential

Electrode potential is defined as "potential difference set up between electrode and electrolyte of same beaker".
It is of two types:

- **Reduction potential**: Tendency of solution to get reduced.
- **Oxidation potential**: Tendency of electrode to get oxidized.

Factors on which electrode potential depends:

- Concentration of ions in solution.
- Nature of metal and its ions.

Electromotive force

It is the potential difference between two electrodes when no current flows through the circuit.

Points to remember

- A chemical reaction may be termed as redox reaction in case there is a change in the oxidation number (increase as well as decrease) of some reacting species which are involved in the reaction.
- The disproportionation reactions are also known as auto-oxidation reactions.
- In the acidic medium first balance the total number of O atoms by adding required number of H_20 molecules to the side deficient in O atoms. Then balance the H atoms by adding required number of H^+ to the side deficient in H atoms.
- In the basic medium first balance the total number of O atoms by adding required number of H_20 molecules to the side deficient in O atoms. Then balance the negative charges by adding required number of OH^- ions on the side deficient in the magnitude of the charges. At the same time, add same number of H_20 molecules on the other side in order to balance the OH^- ions added.
- In general, metals and negative E° values can liberate or evolve hydrogen on reacting with dilute acids while those with positive E° values cannot do so.
- Oxidising agent (oxidant) is a substance undergoing decrease in oxidation number by the gain of electrons.
- Reducing agent (Reductant) is a substance undergoing increase in oxidation number by the loss of electrons.
- Electrochemical cell - It is a device in which the redox reaction is carried indirectly and the decrease in free energy appears as electrical energy.
- Electrode Potential - It is the potential difference between the electrode and its ions in solution.
- Standard Electrode Potential - It is the potential of an electrode with respect to standard hydrogen electrode.
- EMF of cell is the difference in the electrode potentials of the two electrodes in a cell when no current flows through the cell.
- Electrochemical series or activity series has been formed by arranging the metals in order of increasing standard reduction potential values.
- All the alkali metals and some alkaline earth metals are very good reducing agents and displace H_2 from cold water.

- Less active metal eq: Be, Mg, Fe reacts with steam to produce H_2.
- Many metals are also capable of displacing H from acids.
- Halogens are oxidizing Agents.
- Oxidising power decreases down to group.

Oxidising power:$\boldsymbol{F_2 > Cl_2 > Br_2 > I_2}$

1 Mark Questions

Ques1. In the reaction below, identify the species undergoing oxidation and reduction

$H_2S + Cl_2 \rightarrow 2HCl + S$

Ans. The Oxidation number of Sulphur changes from −2 to 0. So, Sulphur is oxidised. The Oxidation number of Chlorine changes from 0 to −1. So, Chlorine is reduced.

Ques2. Why the following Reaction is an example of oxidation reaction?

$CH_4+2O_2 \rightarrow CO_2+2H_2O$

Ans. Methane (CH_4) is oxidized owing to the addition of oxygen to it.

Ques3. Define oxidation in terms of Electron transfer?

Ans. Oxidation is a process in which loss of Electrons takes place.

Ques4. Define an oxidising agent. Name the best oxidising agent.

Ans. Oxidising agent is a substance which can gain electrons easily. F_2 is the best Oxidising agent.

Ques5. Indicate the oxidising and reducing agents in the following reaction:

2Cu2+ + 4I-→2CuI + I2

Ans. Cu is reduced, so Cu {2+} is the oxidizing agent.

I is oxidized, so I {-} is the reducing agent.

Ques6. Given the standard electrode potentials,

K+/K = –2.93V

Ag+/Ag = 0.80V

Hg2+/Hg = 0.79V

Mg2+/Mg = –2.37V

Cr3+/Cr = –0.74V

Arrange these metals in their increasing order of reducing power.

Ans. The reducing agent is stronger as the electrode potential decreases. Hence, the increasing order of the reducing power of the given metals is as given below:

Ag < Hg < Cr < Mg < K

2 Marks Questions

Ques1. The compound AgF_2 is unstable compound. However, if formed, the compound acts as a very strong oxidising agent. Why?

Ans1. The oxidation no. of Ag in AgF(2) AgF_2 is +2. But, +2 is very unstable oxidation no. of Ag. Hence, when AgF{ 2 }Ag_{F2} is formed, silver accepts an electron and forms Ag^{ + }$Ag+$. This decreases the oxidation no. of Ag from +2 to +1. +1 state is more stable. Therefore, AgF{ 2 }AgF_2 acts as a very strong oxidizing agent.

Ques2. Whenever a reaction between an oxidising agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing agent is in excess and a compound of higher oxidation state is formed if the oxidising agent is in excess. Justify this statement giving three illustrations. Justify the above statement with one example.

Ans. When there is a reaction between reducing agent and oxidizing agent, a compound is formed which has lower oxidation number if the reducing agent is in excess and a compound is formed which has higher oxidation number if the oxidizing agent is in excess.

$P4$ and F{ 2 }$F2$ are reducing and oxidizing agent respectively.

In an excess amount of P{ 4 }$P4$ is reacted with F{ 2 }$F2$, then PF_{ 3 }$PF3$ would be produced, where the oxidation no. of P is +3.

$P4(excess)F2 \rightarrow PF3$

If P_{ 4 }$P4$ is reacted with excess of F_{ 2 }$F2$, then PF_{ 5 }$PF5$ would be produced, where the oxidation no. of P is +5.

$P4+F2(excess) \rightarrow PF5$

Ques3. Answer the following questions:

(i) Which non – metals can show disproportionation reaction?

(ii) Which three metals shows disproportionation reaction?

Ans. One of the reacting elements always has an element that can exist in at least 3 oxidation numbers.

(i) The non – metals which can show disproportionation reactions are P, Cl and S.

(ii) The three metals which can show disproportionation reactions are Mn, Ga and Cu.

Ques4. The compound AgF_2 is an unstable compound. However, if formed, the compound acts as a very strong oxidizing agent. Why?

Ans. The oxidation state of Ag in AgF_2 is +2. But, +2 is an unstable oxidation state of Ag. Therefore, whenever AgF_2 is formed, silver readily accepts an electron to form Ag^+. This helps to bring the oxidation state of Ag down from +2 to a more stable state of +1. As a result, AgF_2 acts as a very strong oxidizing agent.