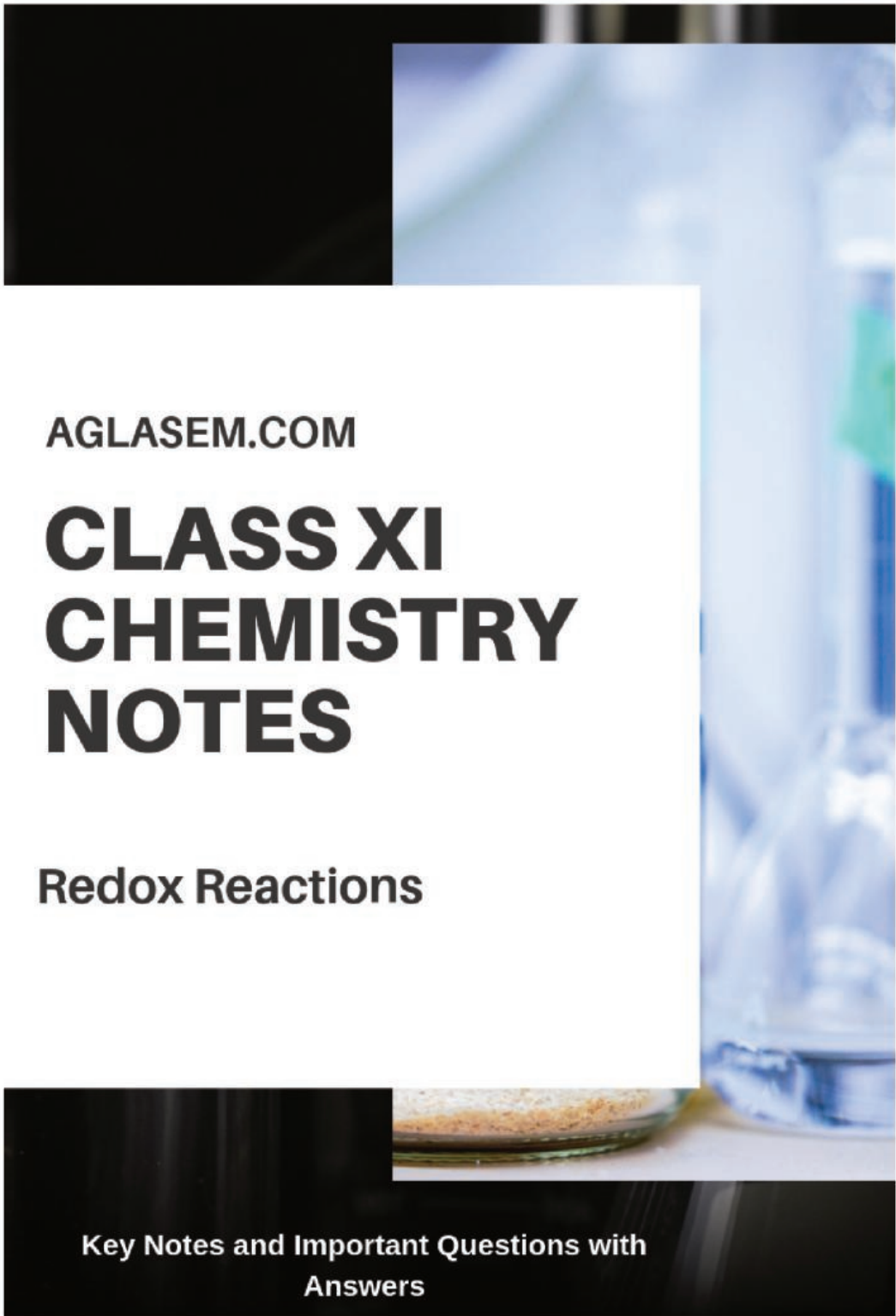


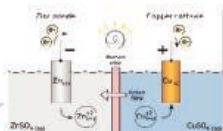
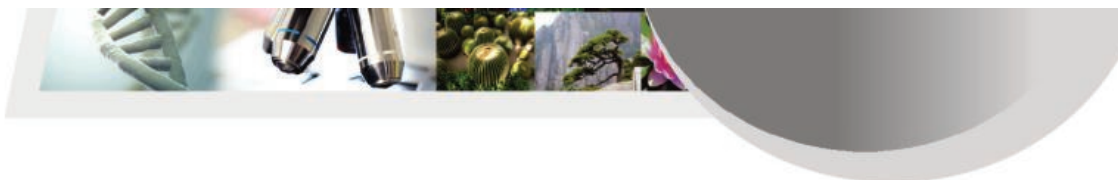
The background of the cover features a photograph of laboratory glassware, including a large Erlenmeyer flask and a graduated cylinder, both containing liquids. The image is partially obscured by a white rectangular area containing the text.

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CLASS XI CHEMISTRY NOTES

Redox Reactions

**Key Notes and Important Questions with
Answers**



Chapter - 8

Redox Reactions

Oxidation and Reduction :

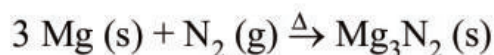
Oxidation	Reduction
1. Addition of oxygen	1. Removal of oxygen
2. Removal an Hydrogen	2. Addition of Hydrogen
3. Addition of an electronegative element.	3. Removal of an electronegative element.
4. Removal of an electropositive element	4. Addition of an electropositive element.
5. Loss of electron(s)	5. Gain of electron(s)
6. Increase in oxidation number.	6. Decrease in oxidation number.

- **Reducing Agent** : Donor of electrons.
- **Oxidising Agent** : Acceptor of electrons.
- **Redox Reaction** : Reactions in which oxidation and reduction takes place simultaneously.
- **Oxidation Number** : It is charge that an atom appears to have in a given species when the bonding electron are counted towards more electro-negative atom.
- **Calculation of Oxidation Number** :
 - Oxidation number of all the elements in their elemental form (in standard state) is taken as zero. Oxidation number of element in a molecule Cl_2 , F_2 , O_2 , P_4 , O_3 , Fe , H_2 , N_2 , C (graphite) is zero.
 - Common Oxidation number of elements of first group is +1. Common Oxidation number of elements of second group + 2.
 - For ions composed of only one atom, the oxidation number is equal to the charge on the ion.

- (d) The oxidation number of oxygen in most compounds is -2 . While in peroxides (e.g., H_2O_2 , Na_2O_2), each oxygen atom is assigned an oxidation number of -1 , in super oxides (e.g., KO_2 , RbO_2) each oxygen atom is assigned an oxidation number of $-\frac{1}{2}$.
- (e) In oxygen difluoride (OF_2) and dioxygen difluoride (O_2F_2), the oxygen is assigned an oxidation number of $+2$ and $+1$, respectively.
- (f) The oxidation number of hydrogen is $+1$ but in metal hydride its oxidation no. is -1 .
- (g) In all its compounds, fluorine has an oxidation number of -1 .
- (h) The algebraic sum of the oxidation number of all the atoms in a compound must be zero.
- (i) In polyatomic ion, the algebraic sum of all the oxidation numbers of atoms of the ion must equal the charge on the ion.

● **Types of Redox Reactions:**

(i) **Combination Reaction :** $0 \quad 0 \quad +2 \quad -3$



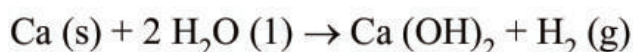
(ii) **Decomposition Reaction :** $+1 \ +5 \ -2 \quad +1 \ -1 \quad 0$



(iii) **Metal Displacement :** $+2 \ +6 \ -2 \quad 0 \quad +2 \ +6 \ -2 \quad 0$

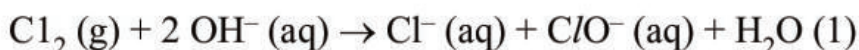


(iv) **Non-metal displacement :** $0 \quad +1 \ -2 \quad +2 \ -2 \ +1 \quad 0$



(v) **Disproportionation reactions :** It is a reaction in which same element is reduced and oxidized simultaneously.

$0 \quad -1 \quad +1$



- **Stock Notation :** Representing oxidation number of metal in Roman numerals within parenthesis after the symbol or name of metal in the molecular formula or name of a compound. For e.g., Stock Notation of Ferric oxide is $\text{Fe}_2\text{(III)O}_3$ or Iron (III) oxide.

- **Fractional Oxidation Number :** When two or more atoms of an element are present in different oxidation states, then calculated oxidation number may comes out as fractional due to average of all the different oxidation states.

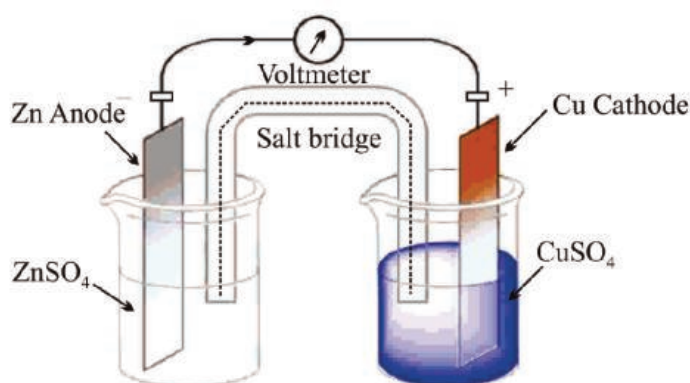
In reality no element can have a fractional oxidation state.

- **Balancing of Redox Reactions :**

(A) Oxidation number method

(B) Half reaction method

- **Electrode Potential (E) :** Potential difference between electrode and electrolytic solution due to charge separation.
- **Standard Electrode Potential (E^0) :** Electrode Potential measured at 298 K and 1M concentration of metal ions (or 1 bar pressure of gas).
- **Electrochemical Cell :** A device in which chemical energy of a spontaneous redox reaction is converted into electrical energy.



Cell diagram: $\text{Zn} \mid \text{Zn}^{2+} \parallel \text{Cu}^{2+} \mid \text{Cu}$

LHS oxidation, $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$

RHS reduction $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$

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

- Representation of an Electrochemical cell :

————— Flow of electrons —————→

←———— Flow of current —————

$\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq}) \parallel \text{Cu}^{2+}(\text{aq}) \mid \text{Cu(s)}$

	Left Electrode	Salt Bridge	Right Electrode
LOAN	Oxidation		Reduction
	Anode		Cathode
	Negative		Positive

- **Functions of Salt Bridge :** (i) To complete inner circuit. (ii) To maintain electrical neutrality around electrodes.

Valency: Valency is the combining capacity of an element. Valency of an element cannot be zero.

Limitation of Concept of Oxidation Number

Acc. to the concept of oxidation number, oxidation means increase oxidation number, during, oxidation there is decrease in electron density while increase in electron density around the atom under going reduction.

1 - Mark Questions

1. Define oxidation and reduction according to electronic concept.
2. Define oxidation and reduction according to oxidation number.
3. A freshly cut apple is almost white but it turns reddish brown after sometime. Give reason.
4. Define oxidation number.
5. Write oxidation number of Mn in KMnO_4 .
6. Write oxidation number of Cr in $\text{Cr}_2\text{O}_7^{2-}$.
7. Write Stock notation of MnO_2 and AuCl_3 .
8. Define redox reaction with example.
9. Define disproportionation reaction. Give one example.
10. Define the term redox titration.
11. Name the indicator used in redox titrations involving $\text{K}_2\text{Cr}_2\text{O}_7$ as an oxidizing agent.
12. At what concentration of Cu^{2+} (aq) will electrode potential become equal to its standard electrode potential ? [Ans. 1 M]
13. The standard reduction potentials of three metals cations X, Y and Z are + 0.52, - 3.03 and - 1.18 V respectively. Arrange X, Y and Z in order of increasing reducing power. [Ans. $X < Z < Y$]
14. An electrochemical cell consists of two electrodes *i.e.*, Anode and Cathode. What is the direction of flow of electrons in this cell ?
15. Why anode is negatively charged in an electrochemical cell.
16. Out of Zn and Cu vessel one will be more suitable to store 1 M HCl ?

[Ans. Cu]

Given $E_{\text{Zn}^{2+}/\text{Zn}}^\theta = -0.76 \text{ V}$, $E_{\text{Cu}^{2+}/\text{Cu}}^\theta = +0.34 \text{ V}$.

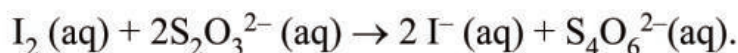
15. Is it safe to stir 1 M AgNO_3 solution with copper spoon ?

Given $E_{\text{Ag}^+/\text{Ag}}^\theta = + 0.80 \text{ V}$, $E_{\text{Cu}^{2+}/\text{Cu}}^\theta = + 0.34 \text{ V}$.

[Ans. No]

2 - Mark Questions

1. Identify oxidant and reductant in the reaction :



2. Calculate oxidation number of Fe in Fe_3O_4 and write a suitable justification of your answer.

3. Oxidation-reduction reactions are complementary. Explain.

4. Write formula for the following compounds :

(i) Mercury (II) chloride

(ii) Nickel (II) sulphate

(iii) Iron (III) sulphate

(iv) Chromium (III) oxide

5. Justify that the reaction : $\text{H}_2\text{O}(\text{s}) + \text{F}_2 \rightarrow \text{HF} + \text{HOF}$ is a redox reaction.

6. A decomposition reaction may or may not be a redox reaction. Write two decomposition reactions in support of the statement.

7. Split the reaction $2 \text{K}(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow 2 \text{KCl}(\text{s})$ into oxidation and reduction half reactions.

8. Calculate the oxidation number of underlined elements in following compounds :

(i) $\text{Ca}\underline{\text{O}}_2$ (ii) $\text{H}_2\underline{\text{S}}_2\text{O}_7$ (iii) $\text{K}_2\underline{\text{Mn}}\text{O}_4$ (iv) $\text{K}\underline{\text{I}}_3$

9. Write the functions of salt bridge in an electrochemical cell.

10. Define the term redox couple. Write the practical application of redox couple.

11. The standard reduction potentials of two metals A and B are $- 0.76 \text{ V}$ and $+ 0.34 \text{ V}$ respectively. An electrochemical cell is formed using electrodes of these metals.

(i) Identify the cathode and anode.

(ii) Write the direction of flow of electron.

3 - Mark Questions

1. Calculate oxidation number of :
 - (i) Cr in $\text{Cr}_2\text{O}_4^{2-}$
 - (ii) O in KO_2
 - (iii) Na in Na_2O_2 .
2. Account for the following :
 - (i) HNO_3 acts as oxidizing agent while HNO_2 can act both as reducing and oxidizing agent.
 - (ii) AgF_2 is unstable compound and act as a strong oxidizing agent.
 - (iii) Ozone acts as an oxidising agent.
3. Permanganate ion (MnO_4^-) reacts with sulphur dioxide gas in acidic medium to produce Mn^{2+} ion and hydrogen sulphate ion. Write ionic equation and balance by ion electron method.
4. Balance the following equation by oxidation number method :
$$\text{P}_4(\text{s}) + \text{OH}^-(\text{aq}) \rightarrow \text{PH}_3 + \text{H}_2\text{PO}_2^-(\text{aq}) \text{ [Basic Medium]}$$
5. Balance the following equation in basic medium :
$$\text{Cl}_2\text{O}_7(\text{g}) + \text{H}_2\text{O}_2(\text{l}) \rightarrow \text{ClO}_2^-(\text{aq}) + \text{O}_2(\text{g})$$
6. Depict the galvanic cell in which the reaction
$$\text{Zn}(\text{s}) + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Ag}(\text{s})$$
takes place. Further show :
 - (i) Which electrode is negatively charged ?
 - (ii) The carriers of the current in the cell
 - (iii) Individual reaction at each electrode.
7. Explain why ?
 - (i) Reaction $\text{FeSO}_4(\text{aq}) + \text{Cu}(\text{s}) \rightarrow \text{CuSO}_4(\text{aq}) + \text{Fe}$ does not occur.
 - (ii) Zinc can displace copper from aqueous CuSO_4 solution but Ag cannot.
 - (iii) Solution of AgNO_3 turns blue when copper rod is immersed in it.

5 - Mark Questions

1. (i) MnO_4^{2-} undergoes disproportionation reaction in acidic medium but MnO_4^- does not. Give reason.
(ii) Give one example each of the following redox reactions:
 - (a) Combination reaction

(b) Decomposition reaction

(c) Metal displacement reaction

2. Consider the cell reaction of an electrochemical cell : $\text{Ni(s)} + 2 \text{Ag}^+(\text{aq}) \rightarrow \text{Ni}^{2+}(\text{aq}) + 2 \text{Ag(s)}$ and answer the following questions :

- (i) Write anode and cathode half reactions.
- (ii) Mention the direction of flow of electrons.
- (iii) How is the electrical neutrality maintained in the solutions of the two half cells.
- (iv) Write the formula for calculating standard emf of this cell.
- (v) How does the emf change when the concentration of silver ions is decreased ?

3. Justify the reason that following reactions are redox reactions.

- (a) $\text{CuO(s)} + \text{H}_2(\text{g}) \longrightarrow \text{Cu(s)} + \text{H}_2\text{O(g)}$
- (b) $\text{Fe}_2\text{O}_3 + 3\text{CO(g)} \longrightarrow 2\text{Fe(g)} + 3\text{CO}_2(\text{g})$
- (c) $\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \longrightarrow 4\text{NO(g)} + 6\text{H}_2\text{O(g)}$
- (d) $\text{BCl}_3(\text{g}) + 3\text{LiAlH}_4 \longrightarrow \text{B}_2\text{H}_6 + \text{LiCl} + \text{AlCl}_3$
- (e) $2\text{K} + \text{F}_2 \longrightarrow 2\text{KF}$

[Hints:– CuO is oxidizing agent, H_2 is acting as reducing agent because Cu (II) is changing to Cu (0) by gain of e^- H_2 is getting oxidised to $\text{H}_2\text{O(g)}$, its oxidation state is changing from 0 to +1, by loss of electrons.

(ii) It is redox reaction: Fe_2O_3 is getting reduced to Fe. CO is getting oxidised to CO_2 .]

4. Using standard electrode : Predict if the reaction between as the following is feasible.

- (i) $\text{Fe}^{3+}(\text{aq})$ and $\text{I}^-(\text{aq})$
- (ii) Ag^+ and Cu
- (iii) Fe^{3+} and $\text{Br}^-(\text{aq})$
- (iv) Ag and $\text{Fe}^{3+}(\text{aq})$
- (iv) $\text{Br}_2(\text{aq})$ and $\text{Fe}^{2+}(\text{aq})$

Hint:– $E^\theta_{\text{I}_2/\text{I}^-} = 0.541 \text{ V}$, $E^\theta_{\text{Cu}^{2+}/\text{Cu}} = 0.34 \text{ V}$, $E^\theta_{\text{Br}_2/\text{Br}^-} = 1.09 \text{ V}$, $E^\theta_{\text{Ag}^+/\text{Ag}} = 0.80 \text{ V}$,
 $E^\theta_{\text{Fe}^{3+}/\text{Fe}^{2+}} = 0.77 \text{ V}$.

5. Draw the diagram for the galvanic cell which would have overall chemical reaction as



Answer the following :

- (i) Write the reactions occurring at each electrode.
- (ii) In which directions do the electrons flow in the external circuit?
- (iii) Name the salt to be taken in salt bridge.
- (iv) Label the anode and cathode.
- (v) How does the EMF change when the concentration of solvers ions is decreased?

