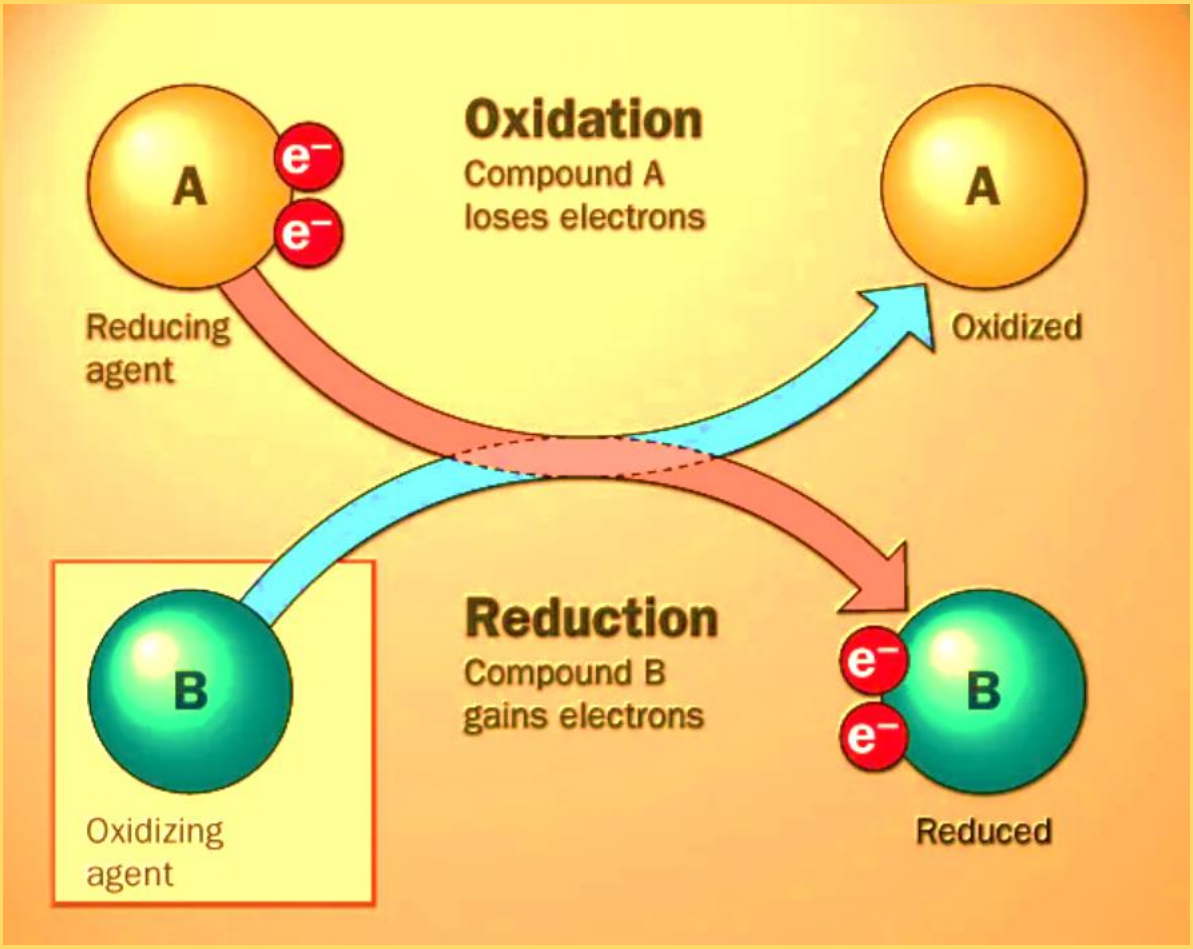
8. REDOX REACTIONS



Chemistry Smart Booklet

Theory + NCERT MCQs +Topicwise Practice

MCQs + NEET PYQs

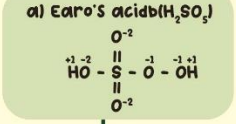
REDOX REACTION

RULES FOR ARRANGING OXIDATION NUMBER (ON)

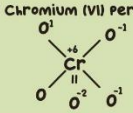
a) Oxidation Number of single element is always 0.
b) Oxidation Number of ions only one atom. ON is equal to charge on ion
c) Oxidation Number of oxygen in most compound is -2
d) Oxidation Number of hydrogen is +1. except when it is bonded to metals in binary compounds.
e) Halogens have an oxidation number of -1. when they occur as halide ions in their compounds.
f) Algebric sum of oxidation number all the atoms in a compounds must be zero

OXIDATION NUMBERS BY STRUCTURE

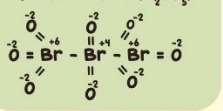
a) Caro's acid (H_2SO_5)



b) Chromium (VI) Peroxide



c) Caro's acid (H_2SO_5)



OXIDATION NUMBER

It indicates the number of electron gained or lost by a particular atom.

STEPS

1. Write the correct formula of the reaction
2. Identify atoms undergoing change in Oxidation Number
3. Calculate i9ncrease or decrease in Oxidation Number per atom amd for entire ion or molecule. If unequal. multiply by suitable number to make equal.
4. Add H+/OH- ion to make total ionic charges of reactants and product equal
5. Equalize H^+ on two sides by adding water.

OXIDATION NUMBER METHOD

1St Method

STEPS

1. Seperate equation into two half reaction.
2. Balance atoms other than O and H
3. For reaction occuring in acidic medium. Add H2O to balance O atoms and H+ to balance H atoms.
4. Balance charges by adding e- to one side of the half reaction.
5. Add two half reactions and cancel the e- on each side.
6. Scale the equation has same type and no. of atoms and same charges on both sides of the equation.

HALF REACTION METHOD

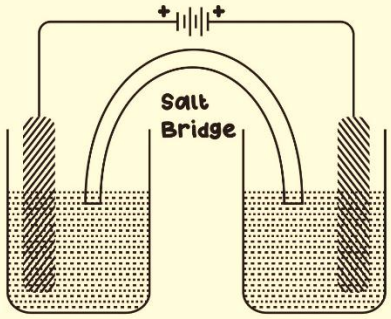
2Nd Method

BALANCING REDOX REACTION

ELECTRO-CHEMICAL SERIES

A series of electrodes on half cells arranged in order of their increasing standard oxidation potentials or in the decreasing order of their standard reduction potential.

GALVENIC CELL



Oxidation half (Anode)

Reduction half (Cathode)

OXIDATION AND REDUCTION

Classical

OXIDATION

Addition of oxygen or removal of hydrogen from a substance.

$C + O_2 \longrightarrow CO_2 + Heat$

Classical

REDUCTION

Removal of oxygen or addition of hydrogen from a substance.

$H_2 + S \longrightarrow H_2S$

Modern

1. Oxidation is loss of electrons.
2. They are considered as reducing agents.
3. Lower oxidation number.

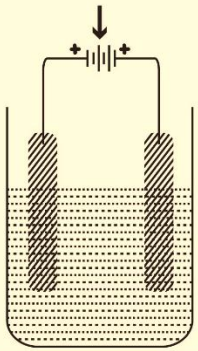
Modern

1. Reduction is gain of electrons.
2. They are considered as oxidising agents.
3. Increases Oxidation Number.

REDOX IN DAILY LIFE

Photosynthesis
Electron of metals
Combination process
Electrochemical cells

APPLICATION



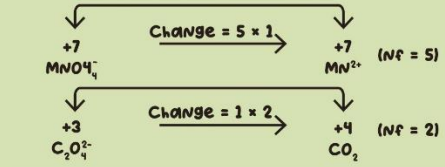
Study of electrode processes and cells

TYPES OF REDOX REACTIONS

COMBINATION REACTION	Two reactants combine to form single products $H_2(g) + O_2(g) \longrightarrow H_2O(l)$
DECOMPOSITION REACTION	Breakdown of a compound into two or more compounds $CaCO_3(S) \xrightarrow{\Delta} CaO(S) + CO_2(g)$
DISPLACEMENT REACTION	An ion/atom in a compound. is replaced by an ion/atom of another elements. $CuSO_4(aq) + Zn(S) \longrightarrow ZnSO_4(aq) + Cu$
DISPROPORTIONATION	An element in one oxidation stateis simultaneously oxidised and reduced $2H_2O_2(l) \longrightarrow 2H_2O(l) + O_2(g)$

CALCULATION OF N- FACTOR

N-factor of oxidising agent/reducing agent = Change in oxidation number per molecule



REDOX REACTIONS

Introduction

Redox reaction is related to gain or loss of electrons. Reaction in which oxidation and reduction takes place simultaneously is called redox reaction. This chapter deals with problems based on redox reactions, oxidation number and balancing of redox reactions by ion, electron method and oxidation number method.

Oxidation Reactions

Oxidation is defined as the addition of oxygen/electronegative element to a substance or rememoval of hydrogen/ electropositive element from a susbtance.

$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$

$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$

Reduction Reactions

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

$2FeCl_3(aq) + H_2(g) \longrightarrow 2FeCl_2(aq) + 2HCl(aq)$

$2HgO(s) \longrightarrow 2Hg(l) + O_2(g)$

Oxidation Number or Oxidation State

Oxidation number for an element is the arbitrary charge present on one atom when all other atoms bonded to it are removed. For example, if we consider a molecule of HCl, the Cl atom is more electronegative than H-atom, therefore, the bonded electrons will go with more electronegative chlorine atom resulting in formation of H^+ and Cl^- ions. So oxidation number of H and Cl in HCl are +1 and –1 respectively.

The following points are important to determine the oxidation number of an element.

1. The oxidation number of an atom in pure elemental form is considered to be zero. e.g., H_2, O_2, Na, Mg.
2. Oxidation number of any element in simple monoatomic ion will be equal to the

charge on that ion, for example, oxidation number of Na in Na^+ is +1.

3. Oxidation number of fluorine in its compound with other elements is always -1.
4. Oxidation number of oxygen is generally -2 but in case of peroxide (H_2O_2) oxygen has oxidation number -1. In a compound OF_2 the oxidation number of oxygen is +2.
5. The oxidation number of alkali metals (Na, K) and alkaline earth metals (Ca, Mg) are +1 and +2 respectively.
6. The oxidation number of halogens is generally -1 when they are bonded to less electronegative elements.
7. Oxidation number of hydrogen is generally +1 in most of its compounds but in case of metal hydride (NaH, CaH_2) the oxidation number is hydrogen is -1.
8. The algebraic sum of the oxidation numbers of all the atoms in a neutral compound is zero. In an ion, the algebraic sum of oxidation number is equal to the charge on that ion.

Oxidising and Reducing Agent

A substance which undergoes oxidation acts as a reducing agent while a substance which undergoes reduction acts as an oxidising agent. For example, we take a redox reaction,

$Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$

In this reaction, Zn is oxidised to Zn^{2+} so Zn is reducing agent and Cu^{2+} is reduced to Cu so Cu^{2+} is an oxidising agent.

Types of Redox Reactions

1. **Combination reactions**

A combination reaction is a reaction in which two or more substances combine to form a single new substance. Combination reactions can also be called synthesis reactions. The general form of a combination reaction is:

$A + B \longrightarrow AB$

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

2. **Decomposition reactions**

A decomposition reaction is a reaction in which a compound breaks down into two or more simpler substances. The general form of a decomposition reaction is:

$AB \longrightarrow A + B$

$2HgO(s) \longrightarrow 2Hg(l) + O_2(g)$

3. **Displacement reactions**

Displacement reaction is a chemical reaction in which a more reactive element displaces a less reactive element from its compound.

$CuSO_4(aq) + Zn(s) \longrightarrow ZnSO_4(aq) + Cu$

4. **Disproportionation reactions**

The reactions in which a single reactant is oxidized and reduced is known as Disproportionation reactions. The disproportionation reaction is given below.

$2H_2O_2 \longrightarrow 2H_2O + O_2$

Balancing of Redox Reactions

a. Oxidation Number Method

In this method number of electrons lost in oxidation must be equal to number of electrons gained in reduction. Following rules are followed for balancing of reactions:

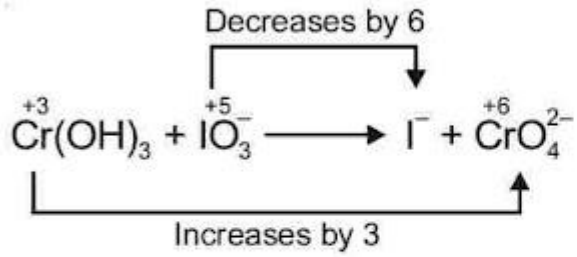
1. Write the skeletal equation of all the reactants and products of the reaction.
2. Indicate the oxidation number of each element and identify the elements undergoing change in oxidation number.
3. Equalize the increase or decrease in oxidation number by multiplying both reactants and products undergoing change in oxidation number by a suitable integer.
4. Balance all atoms other than H and O, then balance O atom by adding water molecules to the side short of O-atoms.
5. In case of ionic reactions:

   i. **For acidic medium:** First balance O atoms by adding H_2O molecules to the side deficient in O atoms and then balance H-atoms by adding H+ ions to the side deficient in H atoms.

   ii. **For basic medium:** First balance O atoms by adding H_2O molecules to whatever side 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.

Balance the ionic equation in alkaline medium

$$Cr(OH)_3 + IO_3^- \longrightarrow I^- + CrO_4^{2-}$$



$$2Cr(OH)_3 + IO_3^- \longrightarrow I^- + 2CrO_4^{2-}$$

Balancing of O and H

$$2Cr(OH)_3 + IO_3^- + 4OH^- \longrightarrow I^- + 2CrO_4^{2-} + H_2O + 4H_2O$$

Now, $2Cr(OH)_3 + IO_3^- + 4OH^- \longrightarrow I^- + 2CrO_4^{2-} + 5H_2O$

b. Ion-Electron Method

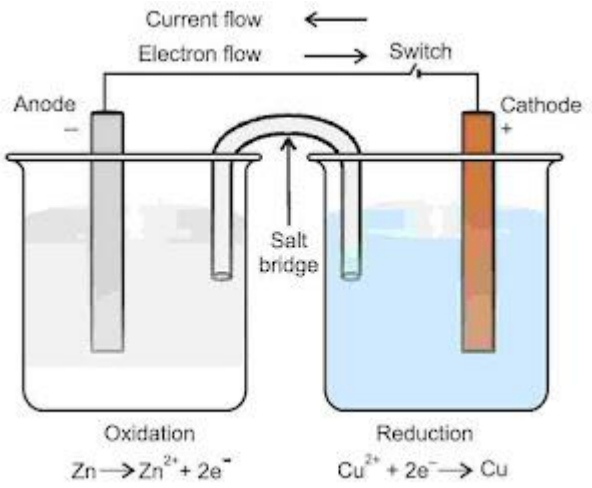
1. Write the skeleton equation and indicate the oxidation number of all the elements which appear in the skeletal equation above their respective symbols.
2. Find out the species which are oxidised and which are reduced.
3. Split the skeleton equation into two half reactions, i.e., oxidation half reaction and reduction half reaction.
4. Balance the two half reaction equations separately by the rules described below:
   - **i.** In each half reaction, 1st 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.
   - **iii.** Balance oxygen atoms by adding required number of H_2O molecules to the side deficient in O atoms.
   - **iv.** 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 OH^- ions are included in the opposite side of the equation.
5. 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.
6. To verify whether the equation thus obtained is balanced or not, the total charge

on either side of the equation must be equal.

Galvanic Cell and Electrode Potential

A galvanic cell or voltaic cell is simple electrochemical cell in which a redox reaction is used to convert chemical energy into electrical energy. It means electricity can be generated with the help of redox reaction in which oxidation and reduction takes place in two separate compartments. Each compartment consists of a metallic conductor and dipped in suitable electrolytic solution of same metal. Metallic rod acts as electrode.

The compartment having electrode dipped in solution of electrolyte is known as half-cell and a half cell has a redox couple. A redox couple means 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 i.e., oxidised form / reduced form. To prepare a galvanic cell two half cells are externally connected through a conducting wire and internally through salt bridge.



Anodic oxidation: $Zn_2 \longrightarrow Zn^{+2}(aq) + 2e(s)$

Cathodic reduction: $Cu^{+2}(aq) + 2e \longrightarrow Cu(s)$

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

This cell can be briefly presented in one line, known as cell notation i.e.,

$Zn \mid Zn^{+2} \mid\mid Cu^{+2} \mid Cu$

Summary-

1. **Oxidation number:** Charge on atom which appears on it when it is present in the combined state.
2. Sum of the oxidation states in a compound/ion should be equal to the zero or to the net charge on the ion.
3. Some elements show variable oxidation states.
4. **Oxidation:** The process in which electrons are lost.
5. **Reduction:** The process in which electrons are gained.
6. **Oxidising agent:** A substance which oxidises the other.
7. **Reducing agent:** A substance which reduces the other.
8. **Redox reaction:** When oxidation and reduction take place together is known as redox reaction.
9. **Disproportionation reaction:** The reaction in which same species is simultaneously oxidised as well as reduced.
10. The change in oxidation state of any element in a compound is useful in calculating the equivalent weight.
11. **Electrochemical series:** Arrangement of E°red of different electrodes in increasing order of electrode potential.
12. **Electrode Potential:** The tendency of an electrode to lose or gain electrons is called electrode potential.
13. The standard electrode potentials of a large number of electrodes have been determined using standard hydrogen electrode as the reference electrode. By convention, the standard electrode potential (E°) of hydrogen electrode is 0.00 volts.

NCERT LINE BY LINE QUESTIONS

(1.) **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, Page: 270]

(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.) Both A and R are false.

(2.) How many electrons appear in the following half reaction when it is balanced) $S_4O_6^{2-} \rightarrow 2S_2O_3^{2-}$ [Page: 276]

(a.) 6
(b.) 2
(c.) 4
(d.) 1

(3.) When does the potential of electrochemical cell become zero [Page: 278]

(a.) E_{ox} of anode and E_{red} of cathode become equal
(b.) E_{red} of anode and E_{ox} of cathode become equal
(c.) E_{red} of anode and E_{ox} of cathode become equal
(d.) Concentration of both the half-cell becomes same

(4.) In the reaction $2H_2S + SO_2 \rightarrow 3S + 2H_2O$ [Page: 264]

(a.) H_2S has been oxidized
(b.) SO_2 has been oxidized
(c.) H_2S is the oxidizing agent
(d.) SO_2 is the oxidizing agent

(5.) When the skeleton oxidation-reduction equation [Page: 275] $MnO_4^- + Br^- \rightarrow Br_2 + Mn^{2+}$ is balanced in acidic solution using the smallest whole number coefficients, the coefficient of Br^- is:

(a.) 8
(b.) 10
(c.) 6
(d.) 4

(6.) **Assertion:** Zinc is a stronger reducing agent than iron.
Reason: Zinc loses electrons more readily than iron. [Page: 266]

(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 nor the correct explanation of A.
(c.) A is true but R is false.
(d.) Both A and R are false.

(7.) Consider the following statement: [page: 272]
(A) $2CuCl \rightarrow CuCl_2 + Cu$ is a disproportionation reaction
(B) All transition metals show disproportionation reaction.
Select the correct option.

(a.) A is true, B is false
(b.) Both A and B are false
(c.) A is false, B is true
(d.) Both A and B are true

(8.) March the items in column I with relevant items in column II. [NCERT Exemplar, Page: 268]

Column I	Column II
(A) Ions having positive charge	(i)+7
(B) The sum of oxidation number of all atoms in a neutral molecule	(ii)-1 (iii)+1
(C) Oxidation number of hydrogen ion $\left(H^{+}\right)$	(iv)0
(D) Oxidation number of fluorine in NaF	(v) Cation
(E) Ions having negative charge	(vi) Anion

(a.) $A \rightarrow (i),\ B \rightarrow (ii),\ C \rightarrow (iii),\ D \rightarrow (iv),\ E \rightarrow (vi)$

(b.) $A \rightarrow (ii),\ B \rightarrow (i)f\ C \rightarrow (ii),\ D \rightarrow (iii),\ E \rightarrow (v)$

(c.) $A \rightarrow (v),\ B \rightarrow (iv),\ C \rightarrow (iii),\ D \rightarrow (ii),\ E \rightarrow (vi)$

(d.) $A \rightarrow (iv),\ B \rightarrow (v),\ C \rightarrow (iii),\ D \rightarrow (i),\ E \rightarrow (ii)$

(9.) Loss of electrons is found in [Page: 265]

(a.) oxidation (b.) reduction

(c.) redox-reaction (d.) All of the above

(10.) Which of the following has been assigned only single oxidation number [Page: 268]

(a.) N (b.) O

(c.) F (d.) H

(11.) Consider the following reaction: [Page: 269] $2MnO_4^- + 5H_2SO_3 \rightarrow 2Mn^{2+} + 3H_2O + 5SO_4^{2-} + 4H^+$ The species that undergoes reduction is

(a.) S (b.) H

(c.) O (d.) Mn

(12.) Consider the following statement: [Page: 279]

(A) Copper liberates hydrogen from a dilute solution of hydrochloric acid.

(B) E of Cu is higher than that of H_2

Select the correct option

(a.) A is true, B is false (b.) Both A and B are false

(c.) A is false, B is true (d.) Both A and B are true

(13.) When the following oxidation-reduction equation representing a reaction that takes place in acidic solution is correctly balanced using the smallest possible whole number coefficients, the coefficients before HSO_3^-, MnO_4^- and H_2O are: [Page: 274] $HSO_3^- + MnO_4^- \rightarrow Mn^{2+} + SO_4^{2-}$

(a.) 4, 2, 2 (b.) 3, 2, 4

(c.) S, 2, 3 (d.) 2, 1, 5

(14.) When copper turnings are added to silver nitrate solution, a blue-coloured solution is formed after some time. It is because copper [Page: 266]

(a.) displaces silver from the solution (b.) forms a blue-coloured complex with $AgNO_3$

(c.) is oxidized to Cu^{2+} (d.) is reduced to Cu^{2+}

(15.) Hydrogen sulphide (H_2S) is a strong reducing agent. Which of the following reactions shows its reducing action? [page: 269]

(a.) $Cd(NO_3)_2 + H_2S \rightarrow CdS + 2HNO_3$ (b.) $CuSO_4 + H_2S \rightarrow CuS + H_2SO_4$

(c.) $2PeCl_3 + H_2S \rightarrow 2FeCl_2 + 2HCI + S$ (d.) $Pb(NO_3)_2 + H_2S \rightarrow PbS + 2CH_3COOH$

(16.) Consider the following statement: [Page: 276]

(A) $\alpha P_4 + bNO_3^- + cH^+ \rightarrow dPO_4^{3-} + eNO_2 + fH_2O$ is a balance reaction in acidic medium.

(B) The ratio of value of b and c in balanced equation is 10: 3.

Select the correct option

(a.) A is true, B is false (b.) Both A and B are false

(c.) A is false, B is true (d.) Both A and B are true

(17.) The oxidation number of Cl in $CaOCl_2$ is/are [Page: 268]

(a.) 0 (b.) +1

(c.) −1 (d.) +1, −1

(18.) Which of the following is a redox reaction? [AIPMT-l997, Page: 264]

(a.) Evaporation of H_2O (b.) Both oxidation and reduction

(c.) H_2SO_4 and NaOH (d.) In atmosphere O_3 from O_2 by lighting

(19.) Which of the following reaction is a redox reaction as well as displacement reaction? [Page: 27l]

(a.) $2HgCl_2 + SnCl_2 \rightarrow HgCl_2 + SnCl_4$ (b.) $ZnO + C \rightarrow Zn + CO$

(c.) $2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$ (d.) $H_2S + Cl_2 \rightarrow 2HCl + S$

(20.) When P reacts with caustic soda, the products are PH_3 and NaH_2P0_2. This reaction is an example of [Page: 267]

(a.) oxidation (b.) reduction

(c.) oxidation and reduction (redox) (d.) neutralization

(21.) The oxidation number of the two nitrogen atoms from left to right in NH_4NO_3 respectively is [Page: 268]

(a.) +1, −1 (b.) −3, +5

(c.) +3, +5 (d.) −1, +1

(22.) Elements whose oxidation number is increased are [Page: 269]

(a.) reduced (b.) hydrolysed

(c.) oxidized (d.) hydrogenated

(23.) Select correct option in reference to e.m.f. series [Page: 278]

(a.) from top to bottom tendency to accept electron increases. (b.) from bottom to top strength as oxidizing agent decreases

(c.) from top to bottom tendency to get oxidized agent increases.

(d.) from bottom to top strength as oxidizing agent increases.

(24.) A standard hydrogen electrode has zero electrode potential because [Page: 277]

(a.) hydrogen is easiest to oxidize.

(b.) this electrode potential is assumed to be zero.

(c.) hydrogen atom has only one electron.

(d.) hydrogen is the lightest element.

(25.) Hot concentrated, sulphuric acid is a moderately strong oxidizing agent. Which of the following reactions does not show oxidizing behaviour [NEET-2O16, Phase-II, Page: 269]

(a.) $Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$

(b.) $3S + 2H_2SO_4 \rightarrow 3SO_2 + 2H_2O$

(c.) $C + 2H_2SO_4 \rightarrow CO_2 + 2SO_2 + 2H_2O$

(d.) $CaP_2 + H_2SO_4 \rightarrow CaSO_4 + 2HP$

(26.) In the balanced state of reaction, $aBr_2 + bOH^- + cH_2O \rightarrow dBrO_4^- + eHBr$ in basic medium. If $c = 9$, then what will be the change in oxidation number? [Page: 276]

(a.) 21

(b.) 7

(c.) 14

(d.) 9

(27.) When the following skeletal equation, which represents a reaction taking place in acid solution, is balanced using smallest possible whole numbers, the coefficient of I^- is:

$Cu^{2+} + I^- \rightarrow CuI + I_2$ [Page: 274]

(a.) 1

(b.) 22

(c.) 3

(d.) 4

(28.) Match the columns: [Page: 279]

Column I	**Column II**
(A) Device in which chemical energy is converted into electrical energy	(i) Electrolysis
(B) Device which keeps electrical neutrality in two half-reactions in an electrochemical cell	(ii) Reduction Reaction
(C) A process in which electric energy is used to bring about decomposition of an electrolyte	(iii) Electrochemical cell
(D) Process which involves reactions between oxidizing and reducing agent to estimate the amount of unknown substance in solution	(iv) Salt bridge

(a.) $A \rightarrow (iii), B \rightarrow (iv), C \rightarrow (i), D \rightarrow (ii)$

(b.) $A \rightarrow (ii), B \rightarrow (i), C \rightarrow (i), D \rightarrow (ii)$

(c.) $A \rightarrow (i), B \rightarrow (i), C \rightarrow (ii), D \rightarrow (ii)$

(d.) $A \rightarrow (ii), B \rightarrow (ii), C \rightarrow (i), D \rightarrow (i)$

(29.) Consider the following statement: [Page: 266]

(A) Silver can be precipitated by adding zinc powder to $AgNO_3$ solution.

(B) $ZnSO_4$ solution can be stirred with a silver spoon.

Select the correct option:

(a.) A is true, B is false

(b.) Both A and B are false

(c.) A is false, B is true

(d.) Both A and B are true

(30.) The oxidation state of Cr in CrO_6 is [Odisha NEET-2019, Page: 268]

(a.) $+4$

(b.) -6

(c.) $+12$

(d.) $+6$

(31.) The oxidation number of P in $H_4P_2O_5, H_4P_2O_6$ and $H_4P_2O_7$ is respectively [Page: 268]

(a.) $+3, +4, +5$

(b.) $+5, +3, +4$

(c.) $+5, +4, +3$

(d.) $+3, +5, +4$

(32.) Match the columns: [Page: 264]

Column I	**Column II**
(a) $2Mg(s)+O_2(g)\rightarrow 2MgO(s)$	(i) removal of hydrogen
(b) $2H_2S(g)+O_2(g)\rightarrow 2S(s)+2H_2O(I)$	(ii) addition of oxygen
(c) $2FeCl_3(aq)+H_2(g)\rightarrow 2PeCl_2(aq)+2HCl(aq)$	(iii) addition of electropositive element
(d) $2HgCl_2(aq)+SnCl_2(aq)\rightarrow Hg_2Cl_2(s)+SnCl_4(aq)$	(iv) removal of electronegative element

(a.) $A\rightarrow(i), B\rightarrow(ii), C\rightarrow(iii), D\rightarrow(iv)$

(b.) $A\rightarrow(ii), B\text{-}\succ(i), C\text{-}*(iv), D\rightarrow(iii)$

(c.) $A\rightarrow(ii), B\rightarrow(iii), C\rightarrow(i), D\rightarrow(iv)$

(d.) $A\rightarrow(iv), B\rightarrow(iii), C\text{-}*(ii), D\text{-}\leftrightarrow(i)$

(33.) 1dentify disproportionation reaction: [NCERT Exemplar Modified, Page: 27O]

(a.) $CH_4+2O_2\rightarrow CO_2+2H_2O$

(b.) $2S_2O_7^{2-}+2H_2O\rightarrow 4SO_4^{-}+4H^{+}$

(c.) $2F_2+2OH^{-}\rightarrow 2F^{-}+OF_2+H_2O$

(d.) $2NO_2+2OH^{-}\rightarrow NO_2^{-}+NO_3^{-}+H_2O$

(34.) A number that is assigned to each atom or ion in a compound which explains its degree of oxidation is called [Page: 267]

(a.) oxidation state

(b.) oxidation number

(c.) oxidation degree

(d.) Both **(a)** and **(b)**

(35.) In the equation [Page: 274] $Cu+xHNO_3\rightarrow Cu(NO_3)_2+yNO_2+2H_2O$ the values of x and y are

(a.) 3 and 5

(b.) 8 and 6

(c.) 4 and 2

(d.) 7 and 1

(36.) **Assertion:** Cu is stronger reducing agent than H_2.

Reason: E^{o} of Cu^{2+}/Cu is negative.

(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.) Both A and R are false.

(37.) The pair of compounds having metals in their highest oxidation state is [Page: 268]

(a.) $MnO_2, PeCl_3$ (b.) MnO_4^-, CrO_2Cl_2

(c.) $\left[Fe(CN)_6\right]^{3-}, \left[CO(CN)_6\right]^{3-}$ (d.) $\left[NiCl_4\right]^{2-}, \left[COCl_4\right]^-$

(38.) Which substance is serving as a reducing agent in the following reaction? [AIPMT-1994, Page: 269] $14H^+ + Cr_2O_7^{2-} + 3Ni \rightarrow 7H_2O + 3Ni^{2+}$

(a.) $H+$ (b.) $Cr_2O_7^{2-}$

(c.) H_2O (d.) Ni

(39.) **Assertion:** The following chemical equation, $2C_6H_6 + 7/20_2 \rightarrow 4CO_2 + 3H_2O$ is a balanced chemical equation.

Reason: In a balanced chemical equation, the total number of atoms of each element may or may not equal on both side of the equation [Page: 276]

(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.) Both A and R are false.

(40.) A mixture of potassium chlorate, oxalic acid and sulphuric acid is heaved. During the reaction which element undergoes maximum change in the oxidation number? [AIPMT-2012, Page: 268]

(a.) S (b.) H

(c.) C1 (d.) C

(41.) The pair of compounds that can exist together is [AIPMT-2O14, Page: 269]

(a.) $FeCl_3, SnCl_2$ (b.) $HgCl_2, SnCl_2$

(c.) $FeCl_2, SnCl_2$ (d.) $FeCl_3$, Kl

(42.) Which of the following cell is different [Page: 278]

(a.) Daniell cell (b.) Lead storage cell

(c.) Leclanche cell (d.) Electrolytic cell

(43.) **Assertion:** F_2 doesnot under go disproportionation reaction.

Reason: Fluorine shows only 0 and–l oxidation state. [Page: 270]

(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 nor the correct explanation of A.

(c.) A is true but R is false. (d.) Both A and R are false.

(44.) Consider the following statement: [Page: 269]

(A) A substance which gets reduced can act as an oxidizing agent.

(B) In the reaction, $3ClO^- \rightarrow ClO_3^- + 2Cl^-$, Cl atom is oxidized as well as reduced.

Select the correct option.

(a.) A is true, B is false (b.) Both A and B are false

(c.) A is false, B is true (d.) Both A and B are true

(45.) A mixture of potassium chlorate, oxalic acid and sulphuric acid is heated. During the reaction which element undergoes maximum change in the oxidation number? [Page: 268]

(a.) S (b.) H

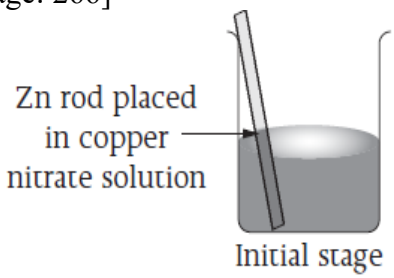
(c.) Cl (d.) C

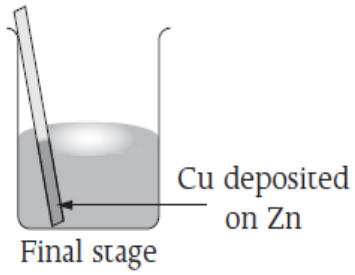
(46.) In an electrochemical cell, which of the following alternative shows $E_{cell} = E_{cell}$? [Page: 27s]

(a.) $k = 1$ (b.) cell reaction is in equilibrium

(c.) concentration of both the half-cell become equal (d.) None of these

(47.) Which of the following redox reactions correctly represents the reaction occurring in a beaker? [Page: 266]





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

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

(c.) $Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + Ag(s)$

(d.) $Cu^{2+}(aq) + Ag(s) \rightarrow Cu(s) + 2Ag^{+}(aq)$

(48.) 2. For $aMg_3N_2 + bH_2O \rightarrow cMg(OH)_2 + dNH_3$, when the equation is balanced, the coefficients a, b, c, d respectively are [page: 274]

(a.) 1, 3, 3, 2 (b.) 1, 6, 3, 2

(c.) 1, 2, 3, 2 (d.) 2, 3, 6, 2

(49.) Sn^{++} loses two electrons in a reaction. What will be the oxidation number of tin after the reaction [Page: 268]

(a.) +2 (b.) Zero

(c.) +4 (d.) –2

(50.) 22. In the following equation: [Page: 276]

$Na_2CO_3 + xHCI \rightarrow 2NaCI + CO_2 + H_2O$, the value of x is

(a.) 1 (b.) 2

(c.) 3 (d.) 4

TOPIC WISE PRACTICE QUESTIONS

TOPIC 1: Oxidation and Reduction Reactions

1. Which quantities are conserved in all oxidation reduction reaction?
   1) Charge only 2) Mass only 3) Both charges and mass 4) Neither charge nor mass
2. In the reaction $H_2S + NO_2 \rightarrow H_2O + NO + S.H_2S$ is
   1) oxidised 2) reduced 3) precipitated 4) None of these
3. The conversion of sugar $C_{12}H_{22}O_{11} \rightarrow CO_2$ is
   1) oxidation 2) reduction
   3) Both oxidation and reduction 4) Neither oxidation nor reduction
4. The product of oxidation of I^- with MnO_4^- in alkaline medium is
   1) IO^- 2) IO_3^- 3) IO_4^- 4) I_2
5. In the following reaction, which is the species being oxidised ?
   $2Fe^{3+}(aq) + 2I^-(aq) \rightarrow I_2(aq) + 2Fe^{2+}(aq)$
   1) Fe^{3+} 2) I^- 3) I_2 4) Fe^{2+}
6. The compound that can work both as an oxidising as well as a reducing agent is
   1) $KMnO_4$ 2) H_2SO_4 3) BaO_2 4) H_2O_2
7. Which of the following substances acts as an oxidising as well as a reducing agent?
   1) Na_2O 2) $SnCl_2$ 3) Na_2O_2 4) $NaNO_2$
8. In the reaction $2FeCl_3 + H_2S \rightarrow 2FeCl_2 + 2HCl + S$
   1) $FeCl_3$ acts as an oxidising agent. 2) Both H_2S and $FeCl_3$ are oxidised.
   3) $FeCl_3$ is oxidised while H_2S is reduced. 4) H_2S acts as an oxidising agent.
9. When iron is rusted, it is
   1) Oxidised 2) reduced 3) evaporated 4) decomposed
10. Which of the following is not an intermolecular redox reaction?
    1) $MgCO_3 \rightarrow MgO + CO_2$ 2) $O_2 + 2H_2 \rightarrow 2H_2O$
    3) $K + H_2O \rightarrow KOH + (1/2)H_2$ 4) $MnBr_3 \rightarrow MnBr_2 + (1/2)Br_2$
11. In the reaction $3Mg + N_2 \rightarrow Mg_3N_2$
    1) magnesium is reduced 2) magnesium is oxidized
    3) nitrogen is oxidized 4) None of these
12. One gas bleaches the colour of flowers by reduction, while the other by oxidation
    1) CO and Cl_2 2) SO_2 and Cl_2 3) H_2S and Br_2 4) NH_3 and SO_2
13. In reaction of $KMnO_4$ and Mohr's salt, $FeSO_4$ is oxidised to
    1) Fe^{2+} 2) Fe^{3+} 3) Fe 4) All of these
14. Match the columns

Column-I	Column-II
1) $2Mg + O_2 \rightarrow 2MgO$	(p) Removal of hydrogen
2) $Mg + Cl_2 \rightarrow MgCl_2$	(q) Removal of electropositive element
3) $2H_2S + O_2 \rightarrow 2S + 2H_2O$	(r) Addition of oxygen
4) $2KI + H_2O + O_3 \rightarrow 2KOH + I_2 + O_2$	(s) Addition of electronegative element, chlorine

   1) A – (s), B – (q), C – (p), D – (r)
   2) A – (r), B – (s), C – (p), D – (q)

3) A – (s), B – (r), C – (q), D – (p)
4) A – (r), B – (p), C – (s), D – (q)

TOPIC 2: Oxidation Number

15. The oxidation number of chromium in potassium dichromate is
1) + 6 2) – 5 3) – 2 4) + 2

16. Phosphorus has the oxidation state of + 3 in
1) phosphorous acid 2) orthophosphoric acid
3) hypophosphorous acid 4) metaphosphoric acid.

17. In which of the following reactions, nitrogen undergoes change in oxidation state?
1) $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$ 2) $2NO_2 \rightarrow N_2O_4$
3) $2NO_2 + H_2O \rightarrow HNO_3 + HNO_2$ d) $N_2O_5 + H_2O \rightarrow 2HNO_3$

18. The brown ring complex compound is formulated as $[Fe(H_2O)_5(NO)]SO_4$. The oxidation state of iron is
1) 1 2) 2 3) 3 4) 0

19. The pair of compounds in which both the metals are in the highest possible oxidation state is
1) $\left[Fe(CN)_4\right]^{4-}, \left[Co(CN)_6\right]^{3-}$ 2) CrO_2Cl_2, MnO_4^-
3) TiO_3, MnO_2 4) $\left[Co(CN)_6\right]^{3-}, MnO_3$

20. The oxidation number of S in $S_2O_8^{2-}$ is
1) +2 2) +4 3) +6 4) +7

21. The oxidation number of oxygen in O_2PtF_6 is
1) –0.5 2) Zero 3) +0.5 4) +1

22. In which of the following compounds, iron has lowest oxidation state?
1) $K_3[Fe(CN)_6]$ 2) $K_4[Fe(CN)_6]$ 3) $FeSO_4.(NH4)_2SO_4.6H_2O$ 4) $Fe(CO)_5$

23. In which of the compounds does 'manganese' exhibit highest oxidation number ?
1) MnO_2 2) Mn_3O_4 3) K_2MnO_4 4) $MnSO_4$

24. On reduction of $KMnO_4$ by oxalic acid in acidic medium, the oxidation number of Mn changes. What is the magnitude of this change?
1) From 7 to 2 2) From 6 to 2 3) From 5 to 2 4) From 7 to 4

25. When SO_2 is passed into an acidified potassium dichromate solution, the oxidation numbers of sulphur and chromium in the final products respectively are
1) +6, +6 2) +6, +3 3) 0, +3 4) +2, +3

26. In which of the following coordination compounds does the transition metal have an oxidation number of +6?
a) $[Cr(H_2O)_5Cl]SO_4$ 2) $Cr\left[\eta^6 - C_6H_6\right]_2$
c) $K_2\left[Cr(CN)_2O_2(O_2)NH_3\right]$ d) $\left[Cr(NH_3)_4(SCN)_2\right]\left[CR(NH_3)_2(SCN)_4\right]$

27. In which of the following transition metal complexes does the metal exhibit zero oxidation state ?
1) $\left[Co(NH_3)_6\right]Cl_3$ 2) $\left[Fe(H_2O)_6\right]SO_4$ 3) $Ni(CO)_4$ 4) $[Fe(H_2O)_6]X_3$

28. Oxidation numbers of chlorine atoms in $CaOCl_2$ are
1) 0, 0 2) –1, –1 3) –1, +1 4) None of these

29. The oxide, which cannot act as a reducing agent, is
1) NO_2 2) SO_2 3) CO_2 4) ClO_2

30. Point out the correct statement of the following about $Na_2S_4O_6$.
1) Average oxidation number of S atom is +2
2) Oxidation number of two S atoms is zero each and that of other two is +5 each
3) Oxidation number of two S atoms is +1 each and that of other two is +4 each
4) None of these

31. The oxidation state of the most electronegative element in the products of the reaction between BaO_2 and H_2SO_4 are
1) 0 and –1 2) –1 and – 2 3) – 2 and 0 4) – 2 and +1

32. When ethane is burnt in presence of excess of oxygen, the oxidation number of carbon changes by
1) +8 2) +7 3) +3 4) +4

TOPIC 3: Disproportionation and Balancing of Redox Reactions

33. Which of the following is not a redox reaction?
1) $4KClO_3 \rightarrow 3KClO_4 + KCl$ 2) $Na_2O + 2HCl \rightarrow 2NaCl + H_2O$
3) $SO_2 + 2H_2S \rightarrow 2H_2O + 3S$ 4) $2Na + O_2 \rightarrow Na_2O_2$

34. For the reaction : $NH_3 + OCl^- \rightarrow N_2H_4 + Cl^-$ in basic medium, the coefficients of NH_3, OCl^- and N_2H_4 for the balanced equation are respectively.
1) 2, 2, 2 2) 2, 2, 1 3) 2, 1, 1 4) 4, 4, 2

35. In the reaction, $Cl_2 + 2OH^- \rightarrow OCl^- + Cl^- + H_2O$
1) OH^- is oxidising and Cl_2 is reducing agent 2) Cl_2 is oxidising and OH^- is reducing agent
3) OH^- is both oxidising and reducing agent 4) Cl_2 is both oxidising and reducing agent

36. When phosphorous reacts with caustic soda, the products are PH_3 and NaH_2PO_2. This reaction is an example of
1) oxidation 2) reduction 3) disproportionation 4) none of these

37. $KMnO_4$ oxidises oxalic acid in acidic medium. The number of CO_2 molecules produced as per the balanced equation is
1) 10 2) 8 3) 6 4) 3

38. Which of the following does not represent redox reaction?
1) $2Ca(OH)_2 + Cl_2 \rightarrow Ca(ClO)_2 + CaCl_2 + 2H_2O$ 2) $Cr_2O_7^{2-} + 2OH^- \rightarrow 2CrO_4^{2-} + H_2O$
3) $NaIO_3 + NaHSO_3 \rightarrow NaHSO_4 + Na_2SO_4 + I_2 + H_2O$ 4) $2Na_2S_2O_3 + I_2 \rightarrow Na_2S_4O_6 + 2NaI$

39. Amongst the following, the strongest oxidising agent is
1) $KMnO_4$ 2) $K_2Cr_2O_7$ 3) H_2O_2 4) O_3

40. Number of moles of $K_2Cr_2O_7$ reduced by one mole of Sn^{2+} ions is
1) 1/3 2) 3 3) 1/6 4) 6

41. The reaction $3ClO^-(aq) \rightarrow ClO_3^-(aq) + 2Cl^-(aq)$ is an example of -
1) Oxidation reaction 2) Reduction reaction
3) Disproportionation reaction 4) Decomposition reaction

42. The species that undergoes disproportionation in an alkaline medium are
1) Cl_2 2) MnO_4^{2-} 3) NO_2 4) All of these

43. What products are expected from the disproportionation reaction of hypochlorous acid?
1) HCl and Cl_2O 2) HCl and $HClO_3$ 3) $HClO_3$ and Cl_2O 4) $HClO_2$ and $HClO_4$

44. Amongst the following which can act as an oxidising as well as reducing agent?
1) F_2 2) SO_3 3) H_2O_2 4) H_2O

45. The number of electrons lost in the following change is $Fe + H_2O \rightarrow Fe_3O_4 + H_2$
1) 2 2) 4 3) 6 4) 8

46. In which of the following reactions, SO_2 behaves as an oxidising agent?

1) $2MnO_4^- + 5SO_2 + 2H_2O \rightarrow 5SO_4^{2-} + 2Mn^{2+} + 4H^+$ 2) $Cl_2 + SO_2 \rightarrow SO_2Cl_2$

3) $2H_2S + SO_2 \rightarrow 3S + 2H_2O$ 4) $SO_2 + H_2O \rightarrow H_2SO_3$

47. The oxidation numbers of C-1, C-2 and C-3 in propyne $\left(\overset{3}{C}H_3\overset{2}{C}\equiv\overset{1}{C}H\right)$ respectively are

1) –1, 0, –3 2) –1, 1, –4 3) –2, 0, –3 4) +1, –2, –3

TOPIC 4: Electrode Potential

48. Standard reduction potentials of the half reactions are given below :

$F_2(g) + 2e^- \rightarrow 2F^-(aq);\ E° = +2.85\ V$

$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq);\ E° = +1.36\ V$

$Br_2(l) + 2e^- \rightarrow 2Br^-(aq);\ E° = +1.06\ V$

$I_2(s) + 2e^- \rightarrow 2I^-(aq);\ E° = +0.53\ V$

The strongest oxidising and reducing agents respectively are :

1) F_2 and I^- 2) Br_2 and Cl^- 3) Cl_2 and Br^- 4) Cl_2 and I_2

49. Standard electrode potentials of redox couples A^{2+}/A, B^{2+}/B, C/C^{2+} and D^{2+}/D are 0.3V, – 0.5V, – 0.75V and 0.9V respectively. Which of these is best oxidising agent and reducing agent respectively –

1) D^{2+}/D and B^{2+}/B 2) B^{2+}/B and D^{2+}/D 3) D^{2+}/D and C^{2+}/C 4) C^{2+}/C and D^{2+}/D

50. The standard electrode potentials of four elements A, B, C and D are –3.05, –1.66, –0.40 and +0.80. The highest chemical reactivity will be exhibited by :

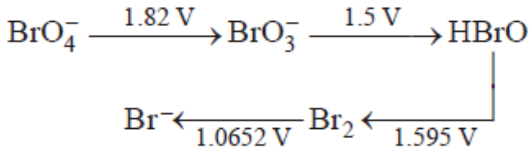
1) A 2) B 3) C 4) D

NEET PREVIOUS YEARS QUESTIONS

1. For the redox reaction $MnO_4^- + C_2O_4^{2-} + H^+ \rightarrow Mn^{2+} + CO_2 + H_2O$ The correct coefficients of the reactants for the balanced equation are [2018]

	MnO_4^-	$C_2O_4^{2-}$	H^+
1)	16	5	2
2)	2	5	16
3)	5	16	2
4)	2	16	5

2. Consider the change in oxidation state of bromine corresponding to different emf values as shown in the diagram below : [2018]



Then the species undergoing disproportionation is

1) BrO_3^- 2) BrO_4^- 3) HBrO 4) Br_2

3. It is because of inability of ns^2 electrons of the valence shell to participate in bonding that : [2017]

1) Sn^{2+} is oxidising while Pb^{4+} is reducing 2) Sn^{2+} and Pb^{2+} are both oxidising and reducing

3) Sn^{4+} is reducing while Pb^{4+} is oxidising 4) Sn^{2+} is reducing while Pb^{4+} is oxidising

4. Assuming complete ionization, same moles of which of the following compounds will require the least amount of acidified $KMnO_4$ for complete oxidation [2015]

1) $FeSO_4$ 2) $FeSO_3$ 3) FeC_2O_4 4) $Fe(NO_2)_2$

5. Nitrogen dioxide and sulphur dioxide have some properties in common. Which property is shown by one of these compounds, but not by the other? [2015]

1) is a reducing agent 2) is soluble in water

3) is used as a food-preservative 4) forms 'acid-rain'

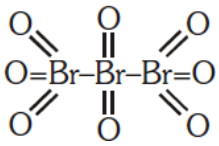
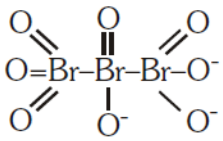
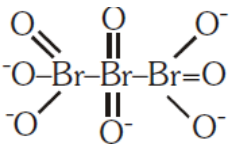
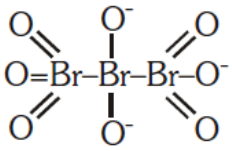
6. The pair of compounds that can exist together is: [2014]

1) $FeCl_3$, $SnCl_2$ 2) $HgCl_2$, $SnCl_2$ 3) $FeCl_2$, $SnCl_2$ 4) $FeCl_3$, KI

7. In acidic medium, H2O2 changes Cr2O7–2 to CrO5 which has two (–O–O) bonds. Oxidation state of Cr in CrO5 is: [2014]

1) + 5 2) + 3 3) + 6 4) – 10

8. The correct structure of tribromooctaoxide is :- [2019]

(1)  (2)  (3)  (4) 

9. The oxidation number of the underlined atom in the following species [2020-COVID-19]
Identify the incorrect option.

(1) $Cu_2\underline{O}$ is -1 (2) $\underline{Cl}O_3^-$ is $+5$ (3) $K_2\underline{Cr_2}O_7$ is $+6$ (4) $H\underline{Au}Cl_4$ is $+3$

10. Which of the following reactions is the metal displacement reaction? Choose the right option

[NEET-2021]

1) $Cr_2O_3 + 2Al \xrightarrow{\Delta} Al_2O_3 + 2Cr$ 2) $Fe + 2HCl \longrightarrow FeCl_2 + H_2\uparrow$

3) $2Pb(NO_3)_2 \longrightarrow 2PbO + 4NO_2 + O_2\uparrow$ 4) $2KClO_3 \xrightarrow{\Delta} 2KCl + 3O_2$

NCERT LINE BY LINE QUESTIONS – ANSWERS

(1.)	a	**(2.)**	b	**(3.)**	b	**(4.)**	a	**(5.)**	b
(6.)	a	**(7.)**	a	**(8.)**	c	**(9.)**	a	**(10.)**	c
(11.)	d	**(12.)**	c	**(13.)**	c	**(14.)**	c	**(15.)**	c
(16.)	d	**(17.)**	d	**(18.)**	b	**(19.)**	c	**(20.)**	c
(21.)	b	**(22.)**	c	**(23.)**	b	**(24.)**	b	**(25.)**	d
(26.)	a	**(27.)**	d	**(28.)**	a	**(29.)**	d	**(30.)**	d
(31.)	a	**(32.)**	b	**(33.)**	d	**(34.)**	d	**(35.)**	c
(36.)	d	**(37.)**	b	**(38.)**	d	**(39.)**	d	**(40.)**	c
(41.)	c	**(42.)**	d	**(43.)**	d	**(44.)**	d	**(45.)**	c
(46.)	d	**(47.)**	a	**(48.)**	b	**(49.)**	c	**(50.)**	b

TOPIC WISE PRACTICE QUESTIONS - ANSWERS

1) 3	**2)** 1	**3)** 1	**4)** 2	**5)** 2	**6)** 4	**7)** 4	**8)** 1	**9)** 1	**10)** 1
11) 2	**12)** 2	**13)** 2	**14)** 2	**15)** 1	**16)** 1	**17)** 3	**18)** 2	**19)** 2	**20)** 3
21) 4	**22)** 4	**23)** 3	**24)** 1	**25)** 2	**26)** 3	**27)** 3	**28)** 3	**29)** 3	**30)** 2
31) 2	**32)** 2	**33)** 2	**34)** 3	**35)** 4	**36)** 3	**37)** 1	**38)** 2	**39)** 4	**40)** 1
41) 3	**42)** 4	**43)** 2	**44)** 3	**45)** 4	**46)** 3	**47)** 1	**48)** 1	**49)** 3	**50)** 1

NEET PREVIOUS YEARS QUESTIONS-ANSWERS

1) 2	**2)** 3	**3)** 4	**4)** 1	**5)** 3	**6)** 3	**7)** 3	**8)** 1	**9)** 1	**10)** 1

NCERT LINE BY LINE QUESTIONS – SOLUTONS

(1.) **(a.)** The decomposition of hydrogen peroxide to form water and oxygen is an example of disproportionation reaction. Because the oxygen of peroxide is in-l oxidation state and it is converted to zero oxidation state in O_2 and -2 oxidation state in H_2O.

(2.) **(b.)** The easiest way to balance these electrons is by balancing the charge. $2- \neq \left(2 \times 2^{-}\right)$, thus adding $2e^{-}$ to the left side will balance the charge. It is a bit trickier to try to use oxidation numbers since the S on the left has an oxidation number of $+2.5$, and the oxidation number of $+2$. This is a change of oxidation state of $+1/2$, four times, thus 2 electrons.

(3.) **(b.)** E_{cell} is zero in equilibrium, that is, E cathode becomes equal to E anode.

(4.) **(a.)** H_2S is changing to S. This is the removal of hydrogen from H_2S. The removal of hydrogen is called oxidation. Thus, H_2S is oxidized to S. Here the substance oxidized is hydrogen sulphide $\left(H_2S\right)$.

(5.) **(b.)** $16H^{+} + 10Br^{-} + 2MnO_4^{-} \rightarrow 5Br_2 + 2Mn^{2+} + 8H_2O$

(6.) **(a.)** Zinc is a stronger reducing agent than iron because it loses electrons more readily.

(7.) **(a.)** Only those transition metals which show three different oxidation state undergo disproportionation reaction.

(8.) (c.)

(9.) **(a.)** According to electronic concept oxidation is defined as a process which involves loss of electrons

(10.) **(c.)** Fluorine (most electronegative) has oxidation number -1 only.

(11.) **(d.)** $MnO_4^{-} \rightarrow Mn^{2+}$

oxidation number changes +7 to +2 in Mn. So, Mn undergoes reduction.

(12.) **(c.)** Copper does not liberate hydrogen from dil.HCl

(13.) **(c.)** $2MnO_4^- + H^+ + 5HSO_3^- \rightarrow 2Mn^{2+} + 3H_2O + 5SO_4^{2-}$

(14.) **(c.)** Cu is above of Ag in electrochemical series and thus

$Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$ reaction occurs.

(15.) **(c.)** In the given reaction H_2 is undergoing oxidation, hence it behaves as reducing agent.

(16.) **(d.)** $aP_4 + 10NO_3^- + 3H^+ \rightarrow dPO_4^{3-} + eNO_2 + fH_2O$

(17.) **(d.)** $CaOCl_2$ is actually $Ca^{2+}(OCl)^{+1}Cl^{-1}$. Therefore, the oxidation number of the two Cl atoms is +l and−l.

(18.) **(b.)** Redox reactions are those chemical reactions which involve transfer of electrons from one chemical species to another.

(19.) **(c.)** This is a redox reaction, where aluminium oxidizes as $Al \rightarrow Al^{3+}$, while hydrogen reduces as $2H^+ \rightarrow H_2$.

This is a displacement reaction also because Al replaces H_2 from HCl.

(20.) **(c.)** $P_4 + 3NaOH + 3H_2O \rightarrow PH_3 + 3NaH_2PO_2$

In reactant P is present in (0) oxidation state and in PH, it is present in (-3) oxidation state and in NaH_2PO_2 it is present in (+l) oxidation state. So, P undergoes both oxidation and reduction.

(21.) **(b.)** NH_4NO_3

NH_4^+ and NO_3^-

Oxidation number of nitrogen in NH_4^+ is $N + 4H = +1 \; N + 4(1) = +1 \; N + 4 = 1 \; N = -3$

Oxidation number of nitrogen in NO_3^- is $N + 3(O) = -1 \; N + 3(-2) = -1 \; N - 6 = -1 \; N = +5$

(22.) **(c.)** Elements whose oxidation number is increased are oxidized.

(23.) **(b.)** More the value of reduction potential, higher is the tendency to accept electron or electrons. Thus, oxidizing nature increases from top to bottom in the electrochemical series. The strength of an oxidizing agent increases as the value of reduction potential becomes more and more positive.

(24.) **(b.)** The electrode potential of a standard hydrogen electrode is arbitrarily assumed to be zero.

(25.) **(d.)** $CaF_2 + H_2SO_4 \rightarrow CaSO_4 + 2HF$

Here, the oxidation state of every atom remains the same so it is not a redox reaction.

(26.) **(a.)** In equation, $aBr_2 + bOH^- + cH_2O \rightarrow dBrO_4^- +$ $eHBr$ in basic medium. If c=9, then the change in oxidation number is 21.

(27.) **(d.)** $Cu^{2+} + 2I^- \rightarrow 2CuI + I_2$

(28.) (a.)

(29.) **(d.)** zinc is more electropositive than silver, when Zinc powder is added into silver nitrate solution Zinc precipitates silver when it reacts with nittrate in the solution.

We can stir $ZnSO_4$ solution with a silver spoon because silver is less reactive then zinc. Thus, it would not displace zinc from the zinc sulphate solution.

(30.) **(d.)** In this compound, the oxygen atoms forms peroxide linkages between them. The oxidation state of oxygen atom in a peroxide is −1.

Let the oxidation state of chromium metal be x. Evaluating the oxidation state, we get. $x + 6(-1) = 0 \; x = 6$

Hence, the oxidation state of chromium in the given compound is +6.

(31.) **(a.)** $H_4P_2O_5 = 4 + 2x - 10 = 0 \Rightarrow -6 + 2x = 0 \Rightarrow 2x = 6 \Rightarrow x = 3$ $H_4P_2O_6 = 4 + 2x - 12 = 0 \Rightarrow 2x - 8 = 0$

$\Rightarrow 2x = 8 \Rightarrow x = 4$ $H_4P_2O_7 = 4 + 2x - 14 = 0 \Rightarrow 2x - 10 = 0$ $\Rightarrow 2x = 10$ $\Rightarrow x = 5$

(32.) (b.)

(33.) **(d.)** Reactions in which the same substance is oxidized as well as reduced are called disproportionation reactions. $2^{+4-2}NO_2 + 2^{-2+1}OH \rightarrow NO_2 + NO_3 + +3-2+5-2H_2O+1-2$

'N' is both oxidized and reduced since the oxidation number of 'N' increases from +4 in NO_2 to $+5$ in NO_3^- and decreases from +4 in NO_2 to +3 in NO_2^-.

(34.) **(d.)** Oxidation state or oxidation number describes the degree of oxidation (loss of electrons) of an atom in a chemical compound.

(35.) **(c.)** $Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$

(36.) **(d.)** E^o of Cu/Cu^{2+} is $+0.34V$ and positive E^o means that the redox couple is a weaker reducing agent than the H^+/H_2 couple.

(37.) **(b.)** The oxidation states of various metals are: (i) $Fe = +3$, $Mn = +4$ (ii) $Cr = +6$, $Mn + 7$ (iii) $Fe = +3$, $Co = +3$ (iv) $Co = +2$, $Mn = +3$

So, the highest oxidation number of Mn in $[MnO_4]^-$ is 7 and Cr in $CrO_2Cl_2 + 6$.

(38.) **(d.)** Since the oxidation number of Ni increases from 0 to 2, it acts as a reducing agent.

(39.) **(d.)** In a balanced chemical equation, the total number of atoms of each element are equal on both sides of the equation.

The correct balanced chemical equation is, $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$

(40.) **(c.)** $2KClO_3 + (COOH) + H_2SO_4 \rightarrow K_2SO_4 + 2CO_2 + 2ClO_2 + 2H_2O$

Maximum change in oxidation number of chlorine, i.e. from +5 to -1.

(41.) **(c.)** Both $FeCl_2$ and $SnCl_2$ are reducing agents with low oxidation numbers.

(42.) **(d.)** Daniell cell, lead storage cell, Leclanche cell are electrochemical cells, which is different than the electrolytic cell.

(43.) **(d.)** F_2 undergoes disproportionation reaction.

F_2 gets reduced to HF and oxidized to HOF.

(44.) **(d.)** Oxidizing agents are electron acceptors and hence get reduced.

(45.) **(c.)** $+1\ +5\ -2\ +6\ +6\ -1$

$KClO_3 + (COOH) + H_2SO_4 \rightarrow K_2SO_4 + KCl + CO_2 + H_2O$

Maximum change in oxidation number of chlorine, i.e. from +5 to −1.

(46.) **(d.)** E_{cell} is the standard state cell potential, which means that the value was determined under standard states. The standard states include a concentration of 1 Molar (mole per litre) and an atmospheric pressure of 1. The E_{cell} is the non-standard state cell potential, which means that it is not determined under a concentration of 1 M and pressure of 1 atm. So, E is E at standard conditions.

(47.) **(a.)** If we place a strip of zinc metal in a copper nitrate solution, we will observe that the oxidation of zinc and reduction of copper takes place. We can write the above reaction as $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

(48.) **(b.)** $Mg_3N_2 + 6H_2O \rightarrow 3Mg(OH)_2 + 2NH_3$

(49.) **(c.)** $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$

(50.) **(b.)** $Na_2CO_3 + 2HCI \rightarrow 2NaCI + CO_2 + H_2O$

TOPIC WISE PRACTICE QUESTIONS – SOLUTIONS

1. (3) Both charge and mass conserves in an oxidation reduction reaction.
2. (1) In this reaction H_2S is oxidised because the oxidation state of 'S' change from – 2 to 0.
3. (1) In this reaction oxidation occur.

4. 2) $6MnO_4^- + I^- + 6OH^- \longrightarrow 6MnO_4^{2-} + IO_3^- + 3H_2O$

5. 2) O.N of I^- is -1and in I_2 O.N. is zero (loss of electrons). Hence I^- oxidised.

6. 4) In H_2O_2 the O.N. of O is –1 which can be increased to 0 or decreased to –2 hence H_2O_2 can work as reducing and oxidising agent.

7. (4) In Na_2O, $SnCl_2$ and Na_2O_2 central atom is either in lowest or highest oxidation state, so it can function either as an oxidising or a reducing agent but not both. However, the oxidation state of N in $NaNO_2$ is +3 which lies between its highest (+5) and lowest (–3) values.

8. 1) In $2FeCl_3 + H_2S \rightarrow 2FeCl_2 + 2HCl + S$ O.N. of S changes from –2 to 0 (hence oxidised) O.N. of Fe changes from +3 to +2 hence reduced.

9. 1) $Fe(s) \rightarrow Fe^{2+} + 2e^-$ (oxidation)

$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ (Reduction)

Atmospheric oxidation:

$$2Fe^{2+} + 2H_2O + \frac{1}{2}O_2 \rightarrow Fe_2O_3 + 4H^+$$

Rest is $Fe_2O_3.xH_2O$

Corrosion or rusting occurs in the presence of water and air.

10. 1) Intermolecular redox reaction and redox reactions are same.

11. (2) In the given reaction, oxidation state of Mg is changing from 0 to +2, while in nitrogen it is changing from 0 to – 3. So oxidation of Mg and reduction of nitrogen takes place.

12. (2) SO_2 bleaches by reduction, while chlorine bleaches colour of flowers by oxidation.

13. 2) $2MnO_4^- + 10Fe^{2+} + 16H^+ \rightarrow 10Fe^{3+} + 2Mn^{2+} + 8H_2O$

14. 2)

15. 1) Let $x =$ oxidation number of Cr in $K_2Cr_2O_7$

$\therefore (2\times1)+(2\times x)+7(-2)=0$ or $2+2x-14=0$ or $x=+6$

16. (1) O.N. of P in H_3PO_3 (phosphorous acid) $3 \times 1 + x + 3 \times (-2) = 0$ or $x = +3$ In orthophosphoric acid (H_3PO_4) O.N. of P is + 5, in hypophosphorous acid (H_3PO_2) it is + 1 while in metaphosphoric acid (HPO_3), it is + 5,

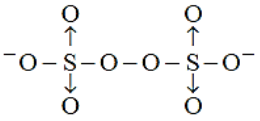
17. (3) O.N. of N in (c) changes from +4 (in NO_2) to +5 (in HNO_3) and +3 in HNO_2.

18. (2) Sum of oxidation state of all atoms in neutral compound is zero.
Let the oxidation state of iron in the complex ion
$[Fe(H_2O)_5(NO)]^{2+}$ SO_4^{2-} be x; then $x + 5 \times 0 + 0 = +2$.
$\therefore\ x = +2$

19. (2) CrO_2Cl_2, MnO_4^-; O.N. of Cr and Mn are +6 and +7 respectively.

20. 3)

$S_2O_8^{2-}$ has the structure :



2 O-atoms in peroxide linkage have –1 O.N. each and rest have –2 each.

21. 4)

22. 4) O.N. of Fe in (a), (b), (c) and (d) respectively are : +3, +2, +2 and 0.

23. 3) O.N. of Mn in K_2MnO_4 is +6

24. (1) In acidic medium MnO_4^- changes to Mn^{2+}, hence O.N. changes from +7 to +2.

25. (2) The redox reaction is : $Cr_2O_7^{2-} + 3SO_2 + 2H^+ \rightarrow 2Cr^{3+} + 3SO_4^{2-} + H_2O$

26. 3)

(a) $x+5x0-1+(-2)=0 \Rightarrow x=+3$

(b) $x+2\times0=0 \Rightarrow x=0$

(c) $2\times1+[x+2(-1)+2(-2)+2(-1)+0]=0 \Rightarrow x=+6$

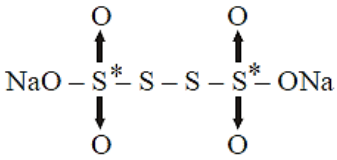
(d) $x+4\times0+2(-1)+x+2\times0+4(-1)=0 \Rightarrow x=+3$

27. 3) In metal carbonyls metal always has O.N. zero.

28. 3) $CaOCl_2$ or $Ca(OCl)$ is the mixed salt of $Ca(OH)_2$ with HCl and HOCl

29. 3) Carbon has the maximum oxidation state of +4, therefore carbon dioxide (CO_2) cannot act as a reducing agent.

30. (2) $Na_2S_4O_6$ has the structure :



O.N. of two S* atoms are +5 each and that of other two S atoms is zero each.

31. 2) $BaO_2+H_2SO_4 \rightarrow BaSO_4+H_2O_2$ The most electronegative element O is present in $BaSO_4$ and H_2O_2, where its O.N. are – 2 and – 1 respectively

32. 2) $C_2H_6+O_2 \rightarrow 2CO_2+3H_2O$

O.N. of carbon changes from –3 (in ethane) to +4 (in CO_2).

33. (2) O.N. of Cl in (a) changes from +5 to +7 and –1. No change of O.N. of any element in (b). In (c), O.N. of S changes from +4 and –2 to zero. In (d), O.N. of Na and O change from 0 each to +1 and –2 respectively.

34. 3) The balanced equation

$2NH_3+OCl^- \rightarrow N_2H_4+Cl^-+H_2O$

35. 4) O.N. of Cl_2 changes from 0 to +1 and -1.

36. 3) O.N. of P changes from zero to –3 in PH_3 and +1 in NaH_2PO_2

37. (1) 10 moles of CO_2 are produced.

38. (2) No element in this reaction has undergone a change in oxidation state.

39. (4) Among these, O_3 has maximum value of standard reduction potential (E° = +2.07V), so it is strongest oxidising agent.

40. 1)

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

$$Sn^{2+}-2e^- \longrightarrow Sn^{4+}$$

On adding

$$\frac{1}{3}Cr_2O_7^{2-}+\frac{14}{3}H^++Sn^{+2} \rightarrow \frac{2}{3}Cr^{+3}+\frac{7}{3}H_2O+Sn^{4+}$$

41. 3) It is a disproportionation reaction because in it an ion is oxidised as well as reduced in a reaction.

42. 4)

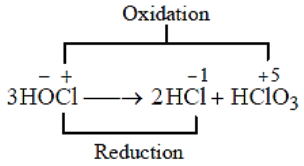
$$Cl_2^0+2NaOH \rightarrow \overset{-1}{NaCl}+\overset{+1}{NaClO}+H_2O$$

$$3\overset{+6}{M}nO_4^{2-}+2H_2O \rightarrow 2\overset{+7}{M}nO_4^-+\overset{+4}{M}nO_2+4OH^-$$

$$2\overset{+4}{N}O_2+H_2O \rightarrow \overset{+5}{H}NO_3+\overset{+3}{H}NO_2$$

All undergo disproportionation

43. 2) During disproportionation same compound undergo simultaneous oxidation reduction.



44. 3) O.N. of O in H_2O_2 is –1. O^- can give up electron to form O_2 or can take up electron to form O^{2-}. Other species act only as oxidising agents.

45. 4)

$$3\overset{0}{Fe} + 8e^- \longrightarrow (Fe^{8/3+})_3$$

$$2H^+ \longrightarrow \overset{0}{H_2} + 2e^-] \times 4$$

$$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$$

Total number of electrons lost or gained is 8.

46. (3) In (a) and (b) SO_2 acts as reducing agent. In (d) it acts neither as reducing nor oxidising agent.

47. (1)

48. (1) Higher the value of reduction potential higher will be the oxidising power whereas the lower the value of reduction potential higher will be the reducing power.

49. (3) The redox couple with maximum reduction potential will be best oxidising agent and with minimum reduction potential will be best reducing agent.

50. (1) Standard electrode potential i.e. reduction potential of A is minimum (–3.05V) i.e. its oxidation potential is maximum which implies 'A' is most reactive chemically.

NEET PREVIOUS YEARS QUESTIONS-EXPLANATIONS

1. 2)

$$\overset{+7}{Mn}O_4^- \longrightarrow \overset{+2}{Mn^{2+}};\ 5e^-\ \text{gain} \qquad ...(i)$$

$$\overset{+3}{C_2}O_4^{2-} \longrightarrow \overset{+4}{CO_2}\ ;\ 2e^-\ \text{loss} \qquad ...(ii)$$

Multiplying (i) by 2 and (ii) by 5 to balance e^-

$$2\,MnO_4^- + 5\,C_2O_4^{2-} \longrightarrow 2\,Mn^{2+} + 10\,CO_2$$

On balancing charge;

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

2. 3) Calculate E°_{cell} corresponding to each compound undergoing disproportionation reaction. The reaction for which E°_{cell} comes out + ve is spontaneous.

$HBrO \rightarrow Br_2$ $\quad E° = 1.595$ V, SRP (cathode)

$HBrO \rightarrow BrO_3^-$ $\quad E° = –1.5$ V, SOP (anode)

$2HBrO \rightarrow Br_2 + BrO_3^-$ $\quad E^\circ_{cell}$ = SRP (cathode) – SRP (anode)

$= 1.595 – 1.5 = 0.095$ V

$E^\circ_{cell\,1} > 0 \Rightarrow \Delta G° < 0$ [spontaneous]

3. 4) Inertness of ns^2 electrons of the valence shell to participate in bonding on moving down the group in heavier p-block elements is called **inert pair effect**. As a result, Pb(II) is more stable than Pb(IV) Sn(IV) is more stable than Sn(II) \ Pb(IV) is easily reduced to Pb(II) and can acts as an oxidising agent whereas Sn(II) is easily oxidised to Sn(IV) and can acts as a reducing agent.

4. 1) Considering the same moles of compounds;

In $Fe^{2+}SO_4^{2-}$ only Fe^{2+} is oxidised by $KMnO_4$.

$\overset{(+6)}{SO_4^{2-}}$ in which sulphur is in highest oxidation state

cannot be oxidised.

In $FeSO_3 \longrightarrow Fe^{2+}$

$\overset{(+4)}{SO_3^{2-}} \longrightarrow \overset{(+6)}{SO_4^{2-}} + 2e^-$

Sulphur can be oxidised to (+6) i.e to SO_4^{2-}

In $FeC_2O_4 \longrightarrow Fe^{2+}$

Carbon can be oxidised to (+IV) i.e to CO_2

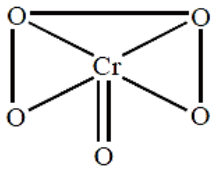
In $Fe(NO_2)_2$

$Fe \longrightarrow Fe^{2+}$

$\overset{(+3)}{2NO_2^-} \rightarrow \underset{(+5)}{2NO_3} + 2e^-$

Nitrogen can be oxidised to (+V) state

5. 3) SO_2 is widely used in food and drinks industries for its property as a preservative and antioxidant while NO_2 is not used as food preservative.
6. (3) Both are reducing agents.
7. (3) The structure of CrO_5 is



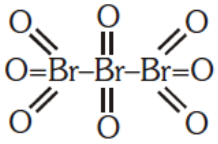
Hence CrO_5 has two peroxy linkage.

Now suppose the ox. no. of Cr is x then

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

$\therefore \quad x = +6$

8. 1)The correct structure is :



other options are anionic

9. 1)Ox. state of "O" in Cu_2O = -2

Ox. state of "Cl" in ClO_3^- = +5

Ox. state of "Cr" in $K_2Cr_2O_7$ = +6

Ox. state of "Au" in $HAuCl_4$ = + 3

$\therefore$ (1) is incorrect

Hoffmann bromamide degradation reaction.

10. 1)A metal in a compound can be displaced by another metal in the uncombined state is metal displacement reaction.