CHAPTER > 08

Redox Reactions

KEY NOTES

- **Oxidation** is the removal of hydrogen or addition of oxygen/electronegative element to a substance or removal electropositive element from a substance.
- **Reduction** is the removal of oxygen/electronegative element from a substance or addition of hydrogen/electropositive element to a substance.
- In terms of electron transfer, loss of electrons or an increase in oxidation state is called **oxidation**.
  Gain of electrons or a decrease in oxidation state is called **reduction**.
- The reactions in which oxidation and reduction occur simultaneously are known as **redox reactions**.
- In the redox reactions, the species which gets oxidised is called **reducing agent**. It is actually the electron donor species.
- In the redox reactions, the species which gets reduced is called **oxidising agent**. It is infact, the electron acceptor species.

Oxidation Number and Oxidants-Reductants

- **Oxidation number** denotes the oxidation state of an element in a compound ascertained according to a set of rules formulated on the basis that electron pair in a covalent bond belongs predominantly to more electronegative element.
- An increase in the oxidation number of the element in the given substance is called **oxidation** and reagents which can increase the oxidation number is called **oxidising agent** or **oxidants**.
- A decrease in the oxidation number of the element in the given substance is called **reduction** and reagents which lowers the oxidation number of an element is called **reducing agent** or **reductants**.
- Some **rules for finding oxidation number** are as follows:
  - Oxidation number of elements in their elementary/free state is 0.
  - For ions composed of only one atom, the oxidation number is equal to the charge on the ion.
  - Fluorine always has –1 oxidation state.
  - Oxidation number of oxygen is –2 (usually). In peroxides it is –1, in superoxides it is $-1/2$ and in OF_2 and O_2F_2, it is $+2$ and + 1, respectively.
  - Oxidation number of H is +1, when combined with non-metals and –1 when combined with metals.

    In a compound, the more electronegative atom will have negative oxidation number, whereas the less electronegative atom will have positive oxidation number.
  - Algebraic sum of oxidation numbers of all the atoms in a neutral molecule is zero and in an ion, it is equal to charge on the ion.
  - For d-block elements,

    oxidation state $= ns$ electrons $+ (n-1)$ d-electrons (unpaired).

    Highest value of oxidation number exhibited by an atom of an element generally increases across the period in the periodic table.

Stock Notation

In Stock notation, the oxidation state of a metal in a compound is expressed by putting a Roman numeral in parenthesis after the symbol of the metal in the molecular formula.

Types of Redox Reactions

Redox reactions can be classified as combination reactions, decomposition reactions, displacement reactions and disproportionation reactions.

- **Combination reactions** are the reactions in which two atoms or molecules combine together to form a third molecule. It may be denoted in the manner,

  $$A + B \longrightarrow C$$

  i.e. $$C + O_2 \longrightarrow CO_2$$
- **Decomposition reactions** are the reactions in which molecule breaks down to form two or more components.

In a decomposition reaction, it is essential that one of the products of decomposition must be in elemental state.

$$AB \longrightarrow A + B$$

i.e. $$2H_2O \longrightarrow 2H_2 + O_2$$

- **Displacement reactions** are the reactions in which an atom (or ion) of a compound is replaced by another ion of same nature.

$$A + BC \longrightarrow AC + B$$

i.e. $$2\,\text{Na}\,(s) + 2H_2O(l) \longrightarrow 2\text{NaOH}(aq) + H_2(s)$$

- **Disproportionation reactions** is a special type of redox reaction in which the same species is simultaneously oxidised as well as reduced.

e.g.

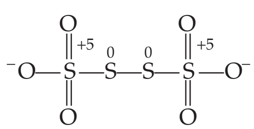
Reduced

$$2\overset{+1}{H_2}\overset{-1}{O_2}(aq) \longrightarrow \overset{+1}{H_2}\overset{-2}{O}(l) + \overset{0}{O_2}(g)$$

Oxidised

Fractional Oxidation State

- The average oxidation state of an element, when two or more of its atoms are present in different oxidation states in a given compound is called **fractional oxidation state.**
- The average oxidation state of the four S-atoms in $S_4O_6^{2-}$ is 2.5, while the actual oxidation states of the four S-atoms are $+5$, 0, 0 and $+5$ in its structure from left to right respectively.



Note *Oxidation number of any element never exceeds its group number.*

Balancing of Redox Reactions

The redox reactions are balanced by two methods. One is **oxidation number method** and second is **half-reaction method** or **ion electron method**.

1. Oxidation Number Method

Step I Write the skeletal equation (if not given, frame it) representing the chemical change.

Step II Assign oxidation numbers to the atoms in the equation and find out which atoms are undergoing oxidation and reduction. Write separate equations for the atoms undergoing oxidation and reduction.

Step III Find the change in oxidation number in each equation. Make the change equal in both the equations by multiplying with suitable integers. Add both the equations.

Step IV Complete the balancing by inspection.

- First balance those substances which have undergone change in oxidation number and then other atoms except hydrogen and oxygen.
- Finally balance hydrogen and oxygen by putting H_2O molecules, wherever needed.
- The final balanced equation should be checked to ensure that there are as many atoms of each element on the right as there are on the left.

Step V In ionic equations, the net charges on both sides of the equation must be exactly the same. Use H^+ ion/ions in acidic reactions and OH^- ion/ions in basic reactions to balance the charge and number of hydrogen and O-atoms.

2. Half-Reaction Method or Ion-Electron Method

Step I Write down the redox reaction in ionic form.

Step II Split the redox reaction into two half-reactions, one for oxidation and other for reduction.

Step III Balance each half-reaction for the number of atoms of each element. For this purpose,

- balance the atoms other than H and O for each half-reaction using simple multiples.
- add water molecules to the side deficient in oxygen and H^+ to the side deficient in hydrogen. This is done in acidic or neutral solutions.
- in alkaline solution, for each deficiency of oxygen, add one water molecule to the same side and $2OH^-$ ions to the other side. If hydrogen is still unbalanced, add one OH^-ion for each excess hydrogen on the same side and one water molecule to the other side.

Step IV Add electrons to the side deficient in electrons to equalise the charge on both sides.

Step V Multiply one or both the half-reactions by a suitable number, so that number of electrons become equal in both the equations.

Step VI Add the two balanced half-reactions and cancel the term common to both sides.

Redox Titrations

In acid-base systems, we come across with a titration method for finding out the strength of one solution against the other using a pH sensitive indicator.

In redox systems, the titration method can be adopted to determine the strength of a reductant or oxidant using a redox sensitive indicator.

The usage of indicators in redox titration is illustrated below :

- The reagent which itself is intensely coloured, e.g. MnO_4^- can acts as the self indicator.
- If there is no dramatic auto-colour change, there are indicators which are oxidised immediately after the last bit of the reactant is consumed.

  e.g. $Cr_2O_7^{2-}$ oxidises the indicator substance diphenylamine just after the equivalence point to produce an intense blue colour.
- Starch is used as indicator in case of reagents which either oxidised I^- (e.g. Cu^{2+}) or reduce I_2 (e.g. $S_2O_3^{2-}$) as it gives intense blue colour with molecular iodine.

Limitations of Concept of Oxidation Number

The concept of redox processes has been evolving with time. In recent past the oxidation process is visualised as decrease in electron density and reduction process as an increase in electron density around the atom(s) involved in the reaction.

Competitive Electron Transfer Reactions

- Competition for electrons between various metals and their ions is done on the basis of electrochemical series. It is a series in which a list of oxidising agents are arranged in decreasing order or reducing agents are arranged in increasing order of their strength.

Release of $2e^-$

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

Gain of $2e^-$

Release of $2e^-$

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

Gain of $2e^-$

Zinc releases electrons to copper and copper releases electrons to silver and, therefore the electron releasing tendency of the metals is in the order :

$$Zn > Cu > Ag .$$

- The negative $E°$ means that, the redox couple is a stronger reducing agent than H^+ / H_2 couple.

  The positive $E°$ means that, the redox couple is a weaker reducing agent than the H^+ / H_2 couple.
- A **redox couple** is defined as the pair having together the oxidised and reduced forms of a substance taking part in an oxidation or reduction half-reaction.
- The potential difference setup between the metal and its own ions in the solution or potential associated with each electrode is called the **electrode potential.**
- If the concentration of each species taking part in the electrode reaction is unity and further the reaction is carried out at 298 K, then the potential of each electrode is said to be **standard electrode potential.**

Mastering NCERT

MULTIPLE CHOICE QUESTIONS

TOPIC 1 ~ *Classical Idea of Redox Reactions*

1 Which of the following processes takes place in oxidation?

(a) Addition of oxygen (b) Addition of hydrogen
(c) Removal of oxygen (d) Removal of chlorine

2 Which of the following processes takes place in reduction?

(a) Removal of oxygen
(b) Addition of hydrogen
(c) Removal of hydrogen
(d) Both (a) and (b)

3 In the following reaction,

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

with respect to Mg, the process is called

(a) oxidation (b) reduction
(c) redox reaction (d) None of these

4 In the given reaction,

$$CH_2{=}CH_2 + H_2 \longrightarrow H_3C{-}CH_3$$

there occurs

(a) oxidation of ethylene
(b) reduction of ethylene
(c) Both (a) and (b)
(d) None of the above

5 In the given reaction,

$$2K_4[Fe(CN)_6](aq) + H_2O_2(aq) \longrightarrow 2K_3[Fe(CN)_6](aq) + 2KOH(aq)$$

Which of the following processes takes place?

(a) Oxidation due to removal of potassium
(b) Oxidation due to removal of iron
(c) Reduction due to removal of potassium
(d) Oxidation due to removal of electronegative element

6 In the given reaction,

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

ferric chloride undergoes

(a) reduction process
(b) oxidation process
(c) addition process
(d) All of the above

7 In the reaction given below, identify the species undergoing oxidation and reduction, respectively

$$H_2S + Cl_2 \longrightarrow 2HCl + S$$

(a) H_2S is oxidised and Cl_2 is reduced
(b) H_2S is reduced and Cl_2 is oxidised
(c) Both H_2S and Cl_2 are oxidised
(d) Both H_2S and Cl_2 are reduced

8 Which of the following reactions represent(s) redox process?

(a) Electrochemical process for extraction of highly reactive metals and non-metals
(b) Manufacturing of caustic soda
(c) Corrosion of metals
(d) All of the above

9 Consider the following reaction,

$$3Fe_3O_4(s) + 8Al(s) \longrightarrow 9Fe(s) + 4Al_2O_3(s)$$

Identify the species undergoing oxidation and reduction, respectively.

(a) Fe_3O_4 is oxidised and Al in reduced
(b) Al is oxidised and Fe_3O_4 is reduced
(c) Both Fe_3O_4 and Al are oxidised
(d) Both Fe_3O_4 and Al are reduced

TOPIC 2 ~ *Redox Reactions in Terms of Electron Transfer Reactions*

10 For the reaction given below,

$$2Na(s) + H_2(g) \longrightarrow 2NaH(s)$$

choose the correct option from the following.

(a) Na is reduced and hydrogen is oxidised
(b) Na is oxidised and hydrogen is reduced
(c) Na undergoes oxidation and hydrogen undergoes reduction
(d) Both (b) and (c)

11 The half-reactions that involve gain of electrons are known as

(a) reduction reactions (b) oxidation reactions
(c) redox reactions (d) All of these

12 An element which donates electrons is known as

(a) reducing agent (b) oxidising agent
(c) complexing agent (d) None of these

13 Sodium sulphide is an ionic compound and written as (by showing its charges)

(a) $(Na^-)_2S^{2-}$ (b) $(Na^+)_2S^{2-}$
(c) $(Na_2^0)S^{2-}$ (d) $(Na^+)_2S^0$

14 The reaction, $2Na + Cl_2 \longrightarrow 2Na^+Cl^-$ involves

(a) loss of $2e^-$ between $2Na \longrightarrow 2Na^+$
(b) gain of $2e^-$ between $Cl_2 \longrightarrow 2Cl^-$
(c) gain of $2e^-$ between $2Na \longrightarrow 2Na^+$
(d) Both (a) and (b)

15 In the reaction, $2Na + Cl_2 \longrightarrow 2Na^+Cl^-$,
the half-reaction, $2Na(s) \longrightarrow 2Na^+(g) + 2e^-$ is called

(a) oxidation half-reaction (b) reduction half-reaction
(c) redox half-reaction (d) None of the above

16 In the reaction, $4Na + O_2 \longrightarrow 2Na_2O$,
sodium acts as a/an

(a) oxidising agent (b) reducing agent
(c) complexing agent (d) None of these

17 In the given reaction, $2Na + S \longrightarrow Na_2S$, sulphur is

(a) oxidised (b) reduced
(c) reducing agent (d) None of these

18 The reaction, $H_2S + I_2 \longrightarrow S + 2HI$, manifests

(a) oxidising nature of I_2 (b) reducing nature of I_2
(c) acidic nature of I_2 (d) alkaline nature of I_2

19 Magnesium reacts with acids producing hydrogen and corresponding magnesium salts. In such reaction, Mg undergoes

(a) reduction (b) oxidation
(c) redox reaction (d) simple dissolution

20 In the reaction,

$$2Na_2S_2O_3 + I_2 \longrightarrow Na_2S_4O_6 + 2NaI,$$

I_2 acts as

(a) oxidising agent
(b) reducing agent
(c) oxidising as well as reducing agent
(d) None of the above

21 In the reaction,

$$2KClO_3 \longrightarrow 2KCl + 3O_2,$$

the elements which have been oxidised and reduced respectively are

(a) chlorine and oxygen
(b) oxygen and chlorine
(c) potassium and oxygen
(d) oxygen and potassium

22 When a strip of metallic zinc is placed in an aqueous solution of copper nitrate, the ions formed is/are

(a) Zn^{2+} (b) Cu^{2+}

(c) NO_2^- (d) Both (a) and (b)

23 In the reaction between copper nitrate solution and zinc, copper ions are reduced by gaining electrons from

(a) copper (b) nitrogen

(c) zinc (d) oxygen

24 Which of the following changes is/are observed. When a clean metallic rod of zinc is placed in a solution of copper nitrate.

(a) Reddish brown copper metal starts depositing on the zinc rod

(b) Zinc rod gradually starts dissolving

(c) Heat is consumed in the reaction

(d) Both (a) and (b)

25 Consider the following reduction half cell,

$$A_2^{x+} + xe^- \longrightarrow A_2$$

Another metal (A_1) which undergoes oxidation and form oxidation half cell. Then, how many electrons are released to complete redox reaction ?

(a) x (b) $x-1$

(c) $x-2$ (d) $1-x$

26 In the given reaction,

Loss of $2e^-$

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

Gain of $2e^-$

Copper metal, $Cu(s)$ is oxidised to $Cu^{2+}(aq)$, while $Ag^+(aq)$ is reduced to silver metal, $Ag(s)$ then the equilibrium greatly lies in favour of

(a) $Ag^+(s)$ and $Cu^{2+}(aq)$ (b) $Ag^+(aq)$ and $Cu^{2+}(s)$

(c) $Ag(s)$ and $Cu^{2+}(aq)$ (d) $Cu^{2+}(aq)$ and $Ag(s)$

27 When a strip of metallic cobalt is placed in aqueous solution of nickel sulphate, the reaction occurs as

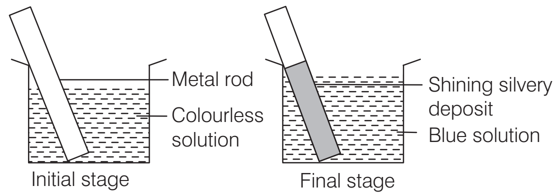
(a) $Co^{2+}(s) + Ni(aq) \longrightarrow Co(aq) + Ni^{2+}(s)$

(b) $Co(s) + Ni^{2+}(aq) \longrightarrow Co(aq) + Ni(s)$

(c) $Co(s) + Ni^{2+}(aq) \longrightarrow Co^{2+}(aq) + Ni(s)$

(d) $Co^{2+}(s) + Ni(aq) \longrightarrow Co(aq) + Ni(s)$

28 Identify the redox reaction taking place in the beaker.



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

(b) $Cu(s) + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$

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

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

29 The correct decreasing order of electron releasing tendency of the given elements is

(a) Ag > Cu > Zn (b) Zn > Cu > Ag

(c) Cu > Ag > Zn (d) Zn > Ag > Cu

TOPIC 3~ Oxidation Number and Its Applications

30 Water molecule is formed by the reaction,

$$2H_2 + O_2 \longrightarrow 2H_2O$$

What does happen in this reaction?

(a) Electrons are transferred from H to O-atom

(b) Electrons are transferred from O to H-atom

(c) Electrons are accepted by H from O-atom

(d) Electrons are donated by O to H-atom

31 In which of the following method, it is assumed that there is a complete transfer of electron from a less electronegative atom to more electronegative atom ?

(a) Reduction number method

(b) Oxidation number method

(c) Redox method

(d) Oxidising agent method

32 An element if present in the free or the uncombined state, its each atom bears an oxidation number of

(a) more than 1 (b) less than 1 (c) more than 2 (d) zero

33 For monoatomic ions, the oxidation number is equal to

(a) charge on the ion (b) zero

(c) more than 1 (d) less than 1

34 In case of peroxides and superoxides, oxidation number of oxygen respectively are

(a) $-1/2$ and -1 (b) -1 and $-1/2$

(c) $+1/2$ and $-1/2$ (d) $+1$ and -1

35 In oxygen difluoride (OF_2) and dioxygen difluoride (O_2F_2), the oxygen is assigned an oxidation number of

(a) +1 and +2 respectively (b) +2 and +2 respectively

(c) +1 and +1 respectively (d) +2 and +1 respectively

36 The algebraic sum of the oxidation number of all the atoms in a compound must be

(a) +1 (b) -1 (c) zero (d) None of these

37 In which of the following compounds, nitrogen exhibits highest oxidation state? ***CBSE AIPMT 2014***

(a) N_2H_4 (b) NH_3 (c) N_3H (d) NH_2OH

38 What is the oxidation number of Cr in $Na_2Cr_2O_7$?

(a) 2 (b) 6 ***JIPMER 2019***

(c) 10 (d) 16

39 The oxidation state of Cr in CrO_5 is ***NEET (Odisha) 2019***

(a) -6 (b) $+12$ (c) $+6$ (d) $+4$

40 Select the compound in which chlorine shows oxidation state of +3?

(a) ClO_4^- (b) ClO_3^- (c) ClO_2^- (d) ClO^-

41 The oxidation number of phosphorus in PCl_5, P_2O_5 and H_2PO_3 respectively are

(a) +5, +2.5 and +4 (b) +5, +5 and +4

(c) +5,+4 and +2.5 (d) +5, +5 and +5

42 The oxidation number of Mn and S in $KMnO_4$ and $Na_2S_2O_3$ respectively are

(a) +7 and +2 (b) +7 and – 2

(c) +7 and +5 (d) +5 and +7

43 Oxidation number of potassium in K_2O, K_2O_2 and KO_2, respectively, is ***JEE Main 2020***

(a) + 1, + 4 and + 2 (b) + 1, + 2 and + 4

(c) + 1, + 1 and + 1 (d) $+2, +1$ and $+\frac{1}{2}$

44 Across a period in the periodic table, the highest value of oxidation number exhibited by an atom of an element from left to right generally

(a) increases

(b) decreases

(c) first increases and then decreases

(d) remains constant

45 The decreasing order of oxidation number of Mn in the given compounds is

I. K_2MnO_4 II. Mn_2O_5 III. MnO_4^- IV. $MnCl_2$

(a) IV > II > III > I (b) IV > III > II > I

(c) III > I > II > IV (d) IV > II > I > III

46 The correct order of N-compounds in its decreasing order of oxidation states is ***NEET 2018***

(a) HNO_3, NH_4Cl, NO, N_2 (b) HNO_3, NO, NH_4Cl, N_2

(c) HNO_3, NO, N_2 NH_4Cl (d) NH_4Cl, N_2, NO, HNO_3

47 Which is the best description behaviour of bromine in the given equation?

$$H_2O + Br_2 \longrightarrow HBr + HOBr$$

(a) Proton acceptor (b) Both oxidised and reduced

(c) Oxidised (d) Reduced

48 What is the average oxidation number of carbon in carbon suboxide which has the following structure, $O{=}C{=}C{=}C{=}O$?

(a) $+\frac{4}{3}$ (b) $+\frac{10}{4}$ (c) +2 (d) $+\frac{2}{3}$

49 The oxidation number of sulphur, chromium and nitrogen in H_2SO_5, $Cr_2O_7^{2-}$ and NO_3^- respectively are

(a) + 8, + 6 and + 5 (b) +6, –6 and +8

(c) + 6, + 6 and + 5 (d) +8, +6 and +7

50 Among which of the following compounds, the oxidation state of nitrogen is positive?

(a) NH_3 (b) HNO_3 (c) Mg_3N_2 (d) NaN_3

51 Oxidation states of *X, Y, Z* are +2, +5 and –2 respectively. The formula of the compound formed by these will be

(a) X_2YZ_6 (b) XY_2Z_6

(c) XY_5 (d) X_3YZ_4

52 Magnesium reacts with an element (X) to form an ionic compound. If the ground state electronic configuration of (X) is $1s^2\ 2s^2\ 2p^3$, the simplest formula for this compound is ***NEET 2018***

(a) Mg_2X (b) MgX_2

(c) Mg_2X_3 (d) Mg_3X_2

53 The average oxidation number of Br in Br_3O_8 is

(a) $\frac{16}{6}$ (b) $\frac{32}{6}$

(c) $\frac{16}{3}$ (d) Both (b) and (c)

54 The oxidation state of a metal in a compound is represented according to the notation which is known as

(a) Alfred stock (b) German stock

(c) Stock notation (d) Haworth stock

55 By using Stock notation, the following compounds, FeO, Fe_2O_3, CuO and MnO_2 can be represented as

(a) Fe(II)O, $Fe_2(III)O_3$, Cu(II)O, $Mn(IV)O_2$, respectively

(b) Fe(III)O, $Fe_2(II)O_3$, Cu(IV)O, $Mn(I)O_2$, respectively

(c) Fe(II)O, $Fe_2(IV)O_3$, Cu(I)O, $Mn(III)O_2$, respectively

(d) Fe(I)O, $Fe_2(I)O_3$, Cu(III)O, $Mn(II)O_2$, respectively

56 An increase in the oxidation number of the element is termed as

(a) reduction (b) oxidation

(c) redox reaction (d) All of these

57 A reagent, which can decrease the oxidation number of an element, is called

(a) reduction (b) oxidant

(c) reducing agent (d) None of these

58 Reactions which involve change in oxidation number of the interacting species is termed as
(a) reducing agent (b) oxidising agent
(c) oxidants (d) redox reactions

59 Identify the species oxidised/reduced in the given reaction,
$$2Cu_2O(s)+Cu_2S(s) \longrightarrow 6Cu(s)+SO_2(g)$$
(a) copper is reduced and sulphur is oxidised
(b) copper is oxidised and sulphur is reduced
(c) oxygen is reduced and copper is oxidised
(d) oxygen is oxidised and sulphur is reduced

60 Which of the following statements is correct about the reaction?
$$2HgCl_2(aq)+SnCl_2(aq) \rightarrow Hg_2Cl_2(s)+SnCl_4(aq)$$
(a) Only oxidation takes place
(b) Only reduction takes place
(c) Both oxidation and reduction takes place
(d) None of the above

61 In the reaction, $2Cu_2O+Cu_2S \longrightarrow 6Cu+SO_2$, sulphur of Cu_2S helps copper both in Cu_2S itself and Cu_2O to decrease its oxidation number, therefore sulphur of Cu_2S is
(a) oxidant (b) reductant
(c) complexing agent (d) None of these

62 In which of the following reactions the underlined substance is oxidised?
(a) $3\,Ca+\underline{N_2} \longrightarrow Ca_3N_2$
(b) $2\,NaI+\underline{Br_2} \longrightarrow 2NaBr+I_2$
(c) $\underline{ZnO}+H_2 \longrightarrow Zn+H_2O$
(d) $\underline{CO}+Cl_2 \longrightarrow COCl_2$

63 When SO_2 is passed through an acidified solution of $KMnO_4$, manganese sulphate is formed. Change in oxidation state of Mn is
(a) +4 to +2 (b) +6 to +3
(c) +7 to +2 (d) None of these

64 Which reaction indicates the action of HNO_3 as oxidising agent?
(a) $NaOH+HNO_3 \longrightarrow NaNO_3+H_2O$
(b) $Ca(OH)_2+2HNO_3 \longrightarrow Ca(NO_3)_2+2H_2O$
(c) $C_6H_6+HNO_3 \longrightarrow C_6H_5NO_2+H_2O$
(d) $NaCl+HNO_3 \longrightarrow HCl+NaNO_3$

65 The correct explanation about the given reaction,
$Cr_2O_7^{2-}+H_2O \longrightarrow 2CrO_4^{2-}+2H^+$ is
(a) chromium is oxidised
(b) chromium is reduced
(c) oxidation number chromium has neither decreased nor increased
(d) hydrogen is reduced

66 The redox reactions are called combustion reactions which make use of
(a) elemental dioxygen (b) elemental dinitrogen
(c) elemental hydrogen (d) metals

67 Identify which of the following reactions is/are combustion reactions?
(a) $C+O_2 \xrightarrow{\Delta} CO_2$
(b) $H_2+Cl_2 \xrightarrow{\Delta} 2HCl$
(c) $CH_4+O_2 \xrightarrow{\Delta} CO_2+H_2O$
(d) Both (a) and (c)

68 A combination redox reaction may be denoted in the manner
$$A+B \longrightarrow C$$
Select the correct option regarding above redox reaction.
(a) Either A or B must be in elemental form
(b) Both A and B must be in elemental form
(c) Both A and B must be in compound form
(d) Both (a) and (b)

69 Which of the following reactions leads to the breakdown of a compound into two or more components?
(a) Combination reactions (b) Displacement reactions
(c) Decomposition reactions (d) None of these

70 The reaction, $2H_2O(l) \xrightarrow{\Delta} 2H_2(g)+O_2(g)$ is an example of
(a) addition reaction (b) decomposition reaction
(c) displacement reaction (d) None of these

71 Which of the following reactions is an example of a redox reaction? ***JEE Main 2017***
(a) $XeF_4+O_2F_2 \longrightarrow XeF_6+O_2$
(b) $XeF_2+PF_5 \longrightarrow [XeF]^+PF_6^-$
(c) $XeF_6+H_2O \longrightarrow XeOF_4+2HF$
(d) $XeF_6+2H_2O \longrightarrow XeO_2F_2+4HF$

72 The decomposition of PCl_5 to PCl_3 and Cl_2 on heating is an example of
(a) intermolecular redox change
(b) intramolecular redox change
(c) intranuclear redox change
(d) None of the above

73 Consider the following decomposition reaction
$$2KClO_3 \longrightarrow 2KCl+3O_2$$
In the above reaction,
(a) potassium is undergoing oxidation
(b) chlorine is undergoing reduction
(c) oxygen is reduced
(d) None of the species are undergoing oxidation or reduction

74 The redox reaction among the following is

(a) reaction of H_2SO_4 with NaOH ***JEE Main 2020***

(b) reaction of $[Co(H_2O)_6]Cl_3$ with $AgNO_3$

(c) combination of dinitrogen with dioxygen at 2000 K

(d) formation of ozone from atmospheric oxygen in the presence of sunlight

75 The given reaction, $CuSO_4 + Zn \longrightarrow Cu + ZnSO_4$ is an example of

(a) metal displacement reaction

(b) non-metal displacement reaction

(c) metal addition reaction

(d) non-metal addition reaction

76 Which of the following is disproportionation reaction? ***JIPMER 2018***

(a) $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$

(b) $CH_4 + 4Cl_2 \longrightarrow CCl_4 + 4HCl$

(c) $2F_2 + 2OH^- \longrightarrow 2F^- + OF_2 + H_2O$

(d) $2NO_2 + 2OH^- \longrightarrow NO_2^- + NO_3^- + H_2O$

77 Which of the following reactions are disproportionation reaction?

I. $2Cu^+ \longrightarrow Cu^{2+} + Cu^0$

II. $3MnO_4^{2-} + 4H^+ \longrightarrow 2MnO_4^- + MnO_2 + 2H_2O$

III. $2KMnO_4 \xrightarrow{\Delta} K_2MnO_4 + MnO_2 + O_2$

IV. $2MnO_4^- + 3Mn^{2+} + 2H_2O \longrightarrow 5MnO_2 + 4H^+$

Select the correct option from the following. ***NEET (National) 2019***

(a) I, II and III (b) I, III and IV

(c) I and IV only (d) I and II only

78 Which of the following will be displaced by all alkali metals from cold water?

(a) Oxygen (b) Hydrogen

(c) Both (a) and (b) (d) None of these

79 The given reactions such as,

I. $Zn + 2\,HCl \longrightarrow ZnCl_2 + H_2$

II. $Fe + 2\,HCl \longrightarrow FeCl_2 + H_2$

are represented as

(a) displacement of zinc and iron metals

(b) displacement of only zinc metals

(c) displacement of only iron metals

(d) displacement of hydrogen

80 When magnesium and iron react with steam, they produce

(a) H_2 (b) O_2 (c) CO_2 (d) None of these

81 Which of the following metals do not react with steam?

(a) Cadmium and magnesium

(b) Cadmium and iron

(c) Tin and cadmium

(d) Magnesium and tin

82 Which of the following metals do not react with HCl?

(a) Cadmium and tin

(b) Silver and gold

(c) Calcium and magnesium

(d) Iron and silicon

83 In the group-17 of the periodic table, the power of elements as oxidising agent decreases from

(a) $F_2 < Cl_2 < Br_2 < I_2$ (b) $I_2 < Br_2 < Cl_2 < F_2$

(c) $I_2 < Cl_2 < Br_2 < F_2$ (d) $F_2 < Cl_2 < I_2 < Br_2$

84 Phosphorus, sulphur and chlorine undergo disproportionation in the

(a) acidic medium (b) alkaline medium

(c) neutral medium (d) Both (a) and (b)

85 Which of the following reactions is responsible for formation of bleaching agent?

(a) $P_4 + 3OH^- + 3H_2O \longrightarrow PH_3 + 3H_2PO_2^-$

(b) $S_8 + 12OH^- \longrightarrow 4S^{2-} + 2S_2O_3^{2-} + 6H_2O$

(c) $Cl_2 + 2OH^- \longrightarrow ClO^- + Cl^- + H_2O$

(d) $H_2SO_4 + P_4 \longrightarrow H_3PO_3 + H_2O$

86 Which of the following halogen does not undergo disproportionation?

(a) Fluorine (b) Bromine

(c) Chlorine (d) Iodine

87 The reaction of white phosphorus with *aq*. NaOH gives phosphine alongwith another phosphorus containing compound. This reaction is known as

(a) simple redox reaction

(b) disproportionation reaction

(c) decomposition reaction

(d) None of the above

88 An example of a disproportionation reaction is ***JEE Main 2019***

(a) $2MnO_4^- + 10\,I^- + 16H^+ \longrightarrow 2Mn^{2+} + 5I_2 + 8H_2O$

(b) $2NaBr + Cl_2 \longrightarrow 2NaCl + Br_2$

(c) $2KMnO_4 \longrightarrow K_2MnO_4 + MnO_2 + O_2$

(d) $2CuBr \longrightarrow CuBr_2 + Cu$

89 Which of the following explanation is incorrect for ClO_4^- species?

(a) ClO_4^- shows disproportionation reaction

(b) ClO_4^- does not show disproportionation reaction

(c) The oxidation number of Cl is +7 in ClO_4^-

(d) ClO_4^- is basic in nature

90 Which of the following halogens react with alkali like other halogens but does not undergo disproportionation reaction?

(a) F_2 (b) Cl_2

(c) Br_2 (d) I_2

91 If a reaction is carried out in acidic medium, then which ions are used to balance the equation?

(a) H^+ ions (b) OH^- ions
(c) H^- ions (d) O^{2-} ions

92 In the given unbalanced reaction,

$Fe^{2+} + Cr_2O_7^{2-} \longrightarrow Fe^{3+} + Cr^{3+}$, which is/are oxidised?

(a) Fe^{2+} (b) $Cr_2O_7^{2-}$
(c) Cr^{3+} (d) All of these

93 In acidic medium, H_2O_2 changes $Cr_2O_7^{2-}$ to CrO_5 which has two (—O—O—) bonds. Oxidation state of Cr in CrO_5 is ***CBSE AIPMT 2014***

(a) + 5 (b) + 3 (c) + 6 (d) – 10

94 Choose the correct explanation regarding half-reaction such as $Cr_2O_7^{2-} \longrightarrow Cr^{3+}$ from the following.

(a) It is oxidation half-reaction
(b) Chromium is being oxidised
(c) $Cr_2O_7^{2-}$ is a good reducing agent
(d) Chromium is being reduced

95 To balance the charges which of the following is added to one side of the half-reaction?

(a) Proton (b) Hydrogen
(c) Oxygen (d) Electrons

96 In the ionic equation,

$BiO_3^- + 6H^+ + xe^- \longrightarrow Bi^{3+} + 3H_2O$, the value of x is

(a) 6 (b) 2 (c) 4 (d) 3

97 Which of the following is correct representation of a given molecular equation in ionic form?

$$6KI + 2KMnO_4 + 4H_2O \longrightarrow 3I_2 + 2MnO_2 + 8KOH$$

(a) $6I^- + 2MnO_4^- + 4H_2O \longrightarrow 3I_2 + 2MnO_2 + 8\overline{O}H$
(b) $6K^+ + 6I^- + 2K^+ + 2MnO_4^- + 4H_2O \longrightarrow 3I_2 + 2MnO_2$
(c) $6K^+ + 6I^- + 2K^+ + 2MnO_4^- + 4H_2O \longrightarrow 3I_2 + 2MnO_2 + 8K^+ + 8\overline{O}H$
(d) $6I^- + 2K^+ + 4H_2O \longrightarrow 3I_2 + 2MnO_2 + 8K^+$

98 $Cu + x\,HNO_3 \longrightarrow Cu(NO_3)_2 + yNO_2 + zH_2O$, the above equation balances, when

(a) $x = 2,\ y = 4,\ z = 3$ (b) $x = 4,\ y = 2,\ z = 2$
(c) $x = 2,\ y = 4,\ z = 2$ (d) $x = 4,\ y = 4,\ z = 2$

99 $a\,K_2Cr_2O_7 + b\,KCl + c\,H_2SO_4 \longrightarrow x\,CrO_2Cl_2 + y\,KHSO_4 + z\,H_2O$,

the above equation balances, when

(a) $a = 2, b = 4, c = 6$ and $x = 2, y = 6, z = 3$
(b) $a = 4, b = 2, c = 6$ and $x = 6, y = 2, z = 3$
(c) $a = 1, b = 4, c = 6$ and $x = 2, y = 6, z = 3$
(d) $a = 1, b = 6, c = 4$ and $x = 6, y = 2, z = 3$

100 For the redox reaction,

$$MnO_4^- + C_2O_4^{2-} + H^+ \longrightarrow Mn^{2+} + CO_2 + H_2O$$

the correct coefficients of the reactants for the balanced equation are ***NEET 2018***

	MnO_4^-	$C_2O_4^{2-}$	H^+
(a)	2	16	5
(b)	2	5	16
(c)	16	5	2
(d)	5	16	2

101 In the following redox reaction,

$xFe^{2+} + yCr_2O_7^{2-} + zH^+ \longrightarrow Fe^{3+} + Cr^{3+} + H_2O$

x, y and z are respectively.

(a) 3, 1 and 14 (b) 6, 1 and 7
(c) 6, 1 and 14 (d) 6, 2 and 14

102 BrO_3^- changes into Br_2 in an acidic medium of a unbalanced equation. How many electron should be present in the balanced equation? ***JIPMER 2019***

(a) 10 electron in left (b) 6 electron in left
(c) 3 electron in left (d) 3 electron in left

103 One mole of acidified $K_2Cr_2O_7$, on reaction with excess KI will liberate n moles of I_2, then the value of n is

(a) 6 (b) 1 (c) 3 (d) 7

104 Which of the following reactions is represented in basic medium?

(a) $MnO_4^-(aq) + SO_2(g) \longrightarrow Mn^{2+}(aq) + HSO_4^-(aq)$
(b) $H_2O_2(aq) + Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + H_2O(l)$
(c) $MnO_4^-(aq) + I^-(aq) \longrightarrow MnO_2(s) + I_2(s)$
(d) $Cr_2O_7^{2-} + SO_2(g) \longrightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$

105 A mixture of potassium chlorate, oxalic acid and sulphuric acid is heated. During the reaction which element undergoes maximum change in oxidation number? ***CBSE AIPMT 2012***

(a) S (b) H
(c) Cl (d) C

106 Consider the following chemical reaction,

$$MnO_4^-(aq) + I^-(aq) \longrightarrow MnO_2(s) + I_2(s)$$

Which of the following reactions is an oxidation half-reaction?

(a) $MnO_4^-(aq) \longrightarrow MnO_2(s)$ (b) $I^-(aq) \longrightarrow I_2(s)$
(c) Both (a) and (b) (d) None of these

107 In redox reaction, the strength of oxidant/ reductant can be determined by

(a) addition method
(b) decomposition method
(c) ion-displacement method
(d) titration method

108 Titration method in acid-base systems is used for finding out the strength of one solution against the other solution by using
(a) hydrogen peroxide indicator
(b) pH sensitive indicator
(c) pOH indicator
(d) None of the above

109 In a solution of $KMnO_4$, MnO_4^- acts as a
(a) induced indicator (b) self indicator
(c) spontaneous indicator (d) None of these

110 $Cr_2O_7^{2-}$ is not a self indicator but
(a) reduces the indicator substance
(b) oxidises the indicator substance
(c) Both (a) and (b) simultaneously
(d) None of the above

111 The reaction of Cu^{2+} ions with iodide ions gives an intense blue colour when starch is added. By which of the following blue colour disappeared?
(a) SO_3^- (b) HSO_4^-
(c) $S_2O_3^-$ (d) SO_2

TOPIC 4 ~ *Redox Reactions and Electrode Processes*

112 When electrons are transferred from Zn to Cu^{2+} in copper sulphate solution, the energy (heat) is
(a) absorbed
(b) evolved
(c) not used in the reaction
(d) None of the above

113 The couple having oxidised and reduced forms of a substance taking part in an oxidation or reduction half-reaction is called
(a) redox couple
(b) oxidised couple
(c) reduced couple
(d) None of the above

114 Redox couple is represented as
(a) Zn^{2+}/Zn (b) Cu^{2+}/Cu
(c) Zn/Zn^{2+} (d) Both (a) and (b)

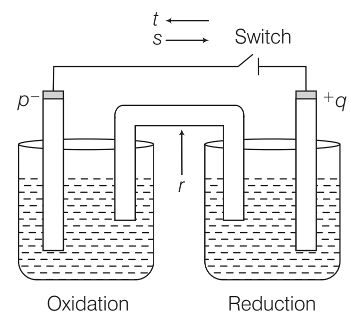
115 In Daniell cell, electrons flow from
(a) cathode to anode
(b) anode to cathode
(c) copper to zinc
(d) SO_4^{2-} to Cu^{2+}

116 In Daniell cell oxidation of zinc and reduction of copper take place at
(a) anode and cathode respectively
(b) cathode and anode respectively
(c) positive electrode and negative electrode respectively
(d) None of the above

117 Solution of potassium chloride or ammonium nitrate in salt-bridge usually solidified by boiling with
(a) agar-agar (b) starch
(c) cellulose (d) glycogen

118 The rods of transition metals such as copper and zinc where potential difference is generated, are termed as
(a) electrodes (b) cathodes
(c) anodes (d) None of these

119 Given below is the set up for Daniell cell, label *p*, *q*, *r*, *s*, t in the given figure.



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

	p	q	r	s	t
(a)	Anode	Cathode	Salt bridge	Electron flow	Current flow
(b)	Cathode	Anode	Salt bridge	Current flow	Electron flow
(c)	Anode	Cathode	Salt bridge	Current flow	Electron flow
(d)	Cathode	Anode	Salt bridge	Ions flow	Electron flow

120 Negative E^S indicates that redox couple is
(a) weaker reducing agent than H^+/H_2 couple
(b) stronger reducing agent than H^+/H_2 couple
(c) stronger oxidising agent than H^+/H_2 couple
(d) None of the above

121 In which of the following reactions H_2O_2 acts as a reducing agent? ***JEE Main 2014***
1. $H_2O_2 + 2H^+ + 2e^- \longrightarrow 2H_2O$
2. $H_2O_2 - 2e^- \longrightarrow O_2 + 2H^+$
3. $H_2O_2 + 2e^- \longrightarrow 2OH^-$
4. $H_2O_2 + 2OH^- - 2e^- \longrightarrow O_2 + 2H_2O$

(a) 1, 2 (b) 3, 4 (c) 1, 3 (d) 2, 4

122 Given,

$$Co^{3+} + e^- \longrightarrow Co^{2+}; E° = +1.81 \text{ V}$$
$$Pb^{4+} + 2e^- \longrightarrow Pb^{2+}; E° = +1.67 \text{ V}$$
$$Ce^{4+} + e^- \longrightarrow Ce^{3+}; E° = +1.61 \text{ V}$$
$$Bi^{3+} + 3e^- \longrightarrow Bi; E° = +0.20 \text{ V}$$

Oxidising power of the species will increase in the order **JEE Main 2019**

(a) $Ce^+ < Pb^{4+} < Bi^{3+} < Co^{3+}$
(b) $Bi^{3+} < Ce^{4+} > Pb^{4+} < Co^{3+}$
(c) $Co^{3+} < Ce^{4+} < Bi^{3+} < Pb^{4+}$
(d) $Co^{3+} < Pb^{4+} < Ce^{4+} < Bi^{3+}$

123 The standard electrode potential (E^-) values of (Al^{3+} / Al, Ag^+ / Ag, K^+ / K and Cr^{3+} / Cr are -1.66 V, 0.80V, 2.93 V and -0.74 V, respectively. The correct decreasing order of reducing power of the metal is **NEET (Odisha) 2019**

(a) Ag > Cr > Al > K (b) K > Al > Cr > Ag
(c) K > Al > Ag > Cr (d) Al > K > Ag > Cr

124 Given, $E°_{Cl_2/Cl^-} = 1.36$ V, $E°_{Cr^{3+}/Cr} = -0.74$ V

$E°_{Cr_2O_7^{2-}/Cr^{3+}} = 1.33$ V, $E°_{MnO_4^-/Mn^{2+}} = 1.51$ V

Among the following, the strongest reducing agent is

(a) Cr (b) Mn^{2+} (c) Cr^{3+} (d) Cl^-

125 Which of the following reaction is not feasible, if the electrode potential are

$E°_{I_2/I^-} = 0.54$V, $E°_{Fe^{3+}/Fe^{2+}} = 0.77$V, $E°_{Ag^+/Ag} = 0.8$V and $E°_{Cu^{2+}/Cu} = 0.34$V

(a) $Fe^{3+}(aq)$ and $I^-(aq)$
(b) $Ag^+(aq)$ and $Cu(s)$
(c) $Fe^{3+}(aq)$ and $Cu(s)$
(d) $Ag(s)$ and $Fe^{3+}(aq)$

SPECIAL TYPES QUESTIONS

I. Statement Based Questions

126 Which of the following statement is incorrect about the elements Cs, Ne, I and F?

(a) F exhibits only negative oxidation state
(b) Cs exhibits only positive oxidation state
(c) Ne exhibits both negative and positive oxidation state
(d) I exhibits both negative and positive oxidation state

127 Identify the correct statement in relation to the following reaction.

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

(a) Zinc is acting as an oxidant
(b) Chlorine is acting as a reductant
(c) Hydrogen ion is acting as a reductant
(d) Zinc is acting as a reductant

128 Which of the following statement is correct regarding the below reaction,

$$3N_2H_4 + 2BrO_3^- \longrightarrow 3N_2 + 2Br^- + 6H_2O$$

(a) N_2H_4 is oxidised and BrO_3^- acts as oxidising agent
(b) BrO_3^- is oxidised and acts as reducing agent
(c) BrO_3^- oxidised and N_2H_4 acts as a oxidising agent
(d) N_2H_4 is reduced and act as oxidising agent

129 Which of the following statement is incorrect?

(a) The reactants, which undergo oxidation and reduction are called reductant and oxidant respectively
(b) In redox reaction, the oxidation number of oxidant increases, while that of reductant decreases
(c) HNO_2 acts as an oxidising as well as reducing agent
(d) Oxidation is the process, in which electrons are lost

130 Consider the galvanic cell reaction,

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

Which of the following statement is incorrect?

(a) Zinc electrode is positively charged
(b) The ions carry current in the cell
(c) At anode, $Zn \longrightarrow Zn^{2+} + 2e^-$
(d) At cathode, $Ag^+(aq) + e^- \longrightarrow Ag(s)$

131 Consider the following redox couples,

$$E°_{Cu^{2+}/Cu} = +0.34 \text{ V},$$
$$E°_{Hg^{2+}/Hg} = +0.885 \text{ V},$$
$$E°_{Ag^+/Ag} = +0.80 \text{ V}$$

Which of the following statement is correct regarding above values?

(a) Mercury can only displace copper from its salt solution
(b) Mercury can displace silver and copper from their salt solution
(c) Copper can only displace silver from its salt solution
(d) Copper can displace mercury and silver from its salt solution

132 Which of the following statements is/are correct about reduction reaction?

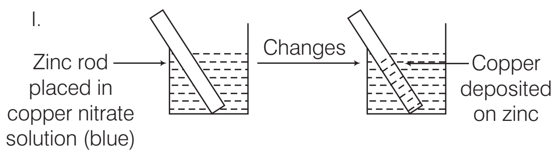
I. Removal of electronegative element.
II. Removal of electropositive element.
III. Addition of electronegative element.
IV. Addition of electropositive element.

Select the correct option.

(a) II and III (b) I and IV
(c) Only I (d) Only III

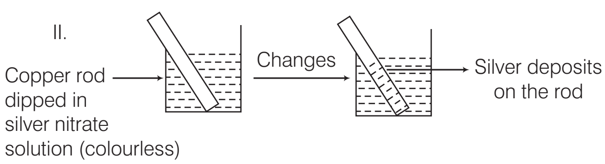
133 Consider the following diagrams and their respective reactions.

I.



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

II.



Reaction :
$Cu(s) + 2Ag^{+}(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$

Which of the above reactions with its respective reactions is/are correct?

(a) Only I (b) Only II
(c) Both I and II (d) Neither I nor II

134 When chlorine gas passes through a concentrated solution of alkali.

I. Cl_2 acts as reducing agent.
II. Cl_2 is reduced
III. Products formed are $5Cl^-$, ClO_3^- and $3H_2O$.

Which of the above statements is/are correct?

(a) Only II (b) I and II
(c) I and III (d) II and III

135 In the given reaction,

$$FeSO_4 + KClO_3 \longrightarrow KCl + Fe_2(SO_4)_3$$

I. $FeSO_4$ acts as reducing agent.
II. $KClO_3$ is reduced.
III. The change in oxidation number of Fe is 2.
IV. The change in oxidation number of Cl is −6.

Which of the above statements is/are correct?

(a) I and III (b) II and III
(c) I, II and IV (d) All of these

136 Consider the following statements about the state of equilibrium for the given reactions.

I. The reaction between Zn and $Cu(NO_3)_2$ solution is product favouring.
II. The reaction between Cu and $AgNO_3$ solution is reactant favouring.
III. The reaction between Co and $NiSO_4$ solution is reactant favouring.

Choose the correct statement(s).

(a) Only I (b) I and II
(c) I and III (d) All of these

137 Consider the following redox couples :

I. $Zn^{2+}/Zn,\ E° = -0.76$
II. $Ag^{2+}/Ag,\ E° = +0.80$ V
III. $Cu^{2+}/Cu,\ E° = +0.34$ V
IV. $Hg^{2+}/Hg;\ E° = +0.885$

Which of the above will act as anode when connected to standard hydrogen electrode which has $E°$ value given as zero?

(a) Only I (b) Only III
(c) I and II (d) II, III and IV

II. Assertion and Reason

■ **Directions** (Q. Nos. 138-144) *In the following questions, a statement of Assertion* (A) *is followed by a corresponding statement of Reason* (R). *Of the following statements, choose the correct one.*

(a) Both A and R are correct; R is the correct explanation of A.
(b) Both A and R are correct; R is not the correct explanation of A.
(c) A is correct; R is incorrect.
(d) A is incorrect; R is correct.

138 **Assertion** (A) Displacement reactions of all halogens using fluorine are not carried out in aqueous solution.

Reason (R) Fluorine attacks water and displaces oxygen of water.

139 **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.

140 **Assertion** (A) In the species, Br_3O_8 each of two extreme bromine exhibits oxidation state of +6 and the middle bromine of +4.

Reason (R) The average of three oxidation numbers of bromine of the Br_3O_8 is 16/3.

141 **Assertion** (A) In the reaction between potassium permanganate and potassium iodide, permanganate ions act as oxidising agent.

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

142 **Assertion** (A) The electrons are transferred from zinc to copper through the wire which connects the two rods.

Reason (R) Electricity flows through the salt-bridge by migration of ions from one beaker to other.

143 **Assertion** (A) A negative value of E° means that the redox couple is a weaker reducing agent than the H^+/H_2 couple.

Reason (R) A negative E° means that the redox couple is stronger reducing agent than the H^+/H_2.

144 **Assertion** (A) Redox couple is the combination of oxidised and reduced form of a substance involved in an oxidation or reduction half-cell.

Reason (R) In the representation $E^{S}_{Fe^{3+}/Fe^{2+}}$ and $E^{S}_{Cu^{2+}/Cu}$, Fe^{3+}/Fe^{2+} and Cu^{2+}/Cu are redox couples.

III. Matching Type Questions

145 Match the Column I (Reaction with underlined species) with Column II (Type of change shown by underlined species) and choose the correct option from the codes given below.

Column I (Reactions)	Column II (Type of change)
A. $\underline{2Mg} + O_2 \longrightarrow 2MgO$	1. Removal of hydrogen.
B. $\underline{Mg} + Cl_2 \longrightarrow MgCl_2$	2. Removal of electropositive element.
C. $\underline{2H_2S} + O_2 \longrightarrow 2S + 2H_2O$	3. Addition of oxygen.
D. $\underline{2KI} + H_2O + O_3 \longrightarrow 2KOH + I_2 + O_2$	4. Addition of electronegative element.

Codes

	A	B	C	D		A	B	C	D
(a)	2	3	4	1	(b)	3	4	1	2
(c)	3	4	2	1	(d)	3	2	1	4

146 Match the items in Column I with relevant item given in Column II and select the correct option from codes given below.

Column I	Column II
A. Ions having positive charge.	1. +7
B. The sum of oxidation number of all atoms in a neutral molecule.	2. −1
C. Oxidation number of hydrogen ion	3. +1
D. Oxidation number of fluorine in NaF	4. 0
E. Ions having negative charge	5. Cation
	6. Anion

Codes

	A	B	C	D	E		A	B	C	D	E
(a)	5	4	3	2	6	(b)	1	4	3	5	6
(c)	2	1	3	4	5	(d)	6	2	3	4	5

147 Match the Column I with Column II and select the correct option from the codes given below.

Column I (Compounds)	Column II (Oxidation number of nitrogen)
A. N_2O_3	1. − 3
B. HNO_3	2. +5
C. NO	3. +3
D. NH_4OH	4. +2

Codes

	A	B	C	D		A	B	C	D
(a)	2	3	4	1	(b)	1	2	3	4
(c)	4	1	2	3	(d)	3	2	4	1

148 Match the Column I with Column II and select the correct option for oxidation number of N-atom from the codes given below.

Column I (Compounds)	Column II (Oxidation number)
A. NH_2OH	1. − 1
B. Mg_3N_2	2. − 1
C. N_2O	3. +5
D. N_2O_5	4. − 3

Codes

	A	B	C	D		A	B	C	D
(a)	1	3	4	2	(b)	2	4	3	1
(c)	2	4	1	3	(d)	4	2	1	3

149 Match the Column I with Column II and select the correct answer for oxidation number of iron, nickel, nitrogen and xenon from given codes.

Column I	Column II
A. $K_4[\underline{Fe}(CN)_6]$	1. −3, +5
B. $[\underline{Ni}(CN)_4]^{2-}$	2. +8
C. $\underline{N}H_4\underline{N}O_3$	3. +2
D. $Ba_2\underline{Xe}O_6$	4. −2

Codes

	A	B	C	D		A	B	C	D
(a)	3	3	1	2	(b)	3	4	2	1
(c)	2	1	3	4	(d)	1	2	4	3

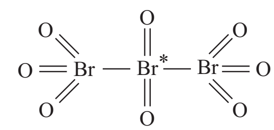
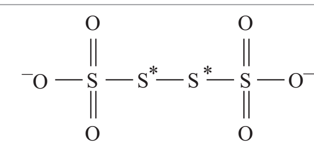
150 Match Column I with Column II and select the correct answer from given codes.

Column I (Types of redox reactions)	Column II (Redox reactions)
A. Combustion reaction	1. $CH_4(g) + 2O_2 \xrightarrow{\Delta} CO_2(g) + 2H_2O(l)$
B. Disproportionation reaction	2. $2NaH(s) \xrightarrow{\Delta} 2Na(s) + H_2(g)$
C. Decomposition reaction	3. $S_8(s) + 12OH^-(aq) \longrightarrow 4S^{2-}(aq) + 2S_2O_3^{2-}(aq) + 6H_2O(l)$
D. Displacement reaction	4. $Cr_2O_3(s) + 2Al(s) \xrightarrow{\Delta} Al_2O_3(s) + 2Cr(s)$

Codes

	A	B	C	D
(a)	4	2	1	3
(b)	1	2	4	3
(c)	1	3	2	4
(d)	2	1	4	3

151 Match the Column I with column II and select the correct option from the codes given below.

	Column I (Species)	**Column II** (Oxidation number of element marked with asterisk)
A.	$O{=}C{=}C{=}\overset{*}{C}{=}O$	1. 0
B.		2. +4
C.		3. +2
		4. +6

Codes:

	A	B	C
(a)	2	3	1
(b)	3	2	1
(c)	1	2	4
(d)	3	1	4

NCERT & NCERT Exemplar

MULTIPLE CHOICE QUESTIONS

NCERT

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

(a) The oxidation state of Ag in AgF_2 is $+2$ which is unstable
(b) The oxidation state of Ag in AgF_2 is $+1$ which is stable
(c) The oxidation state of Ag in AgF_2 is $+1$ which is unstable
(d) The oxidation state of Ag in AgF_2 is $+2$ which is stable

153 Identify the substance oxidised, reduced, oxidising agent and reducing agent respectively from the following reaction.

$$2AgBr(s) + C_6H_6O_2(aq) \longrightarrow 2Ag(s) + 2HBr(aq) + C_6H_4O_2(aq)$$

(a) AgBr, $C_6H_6O_2$, $C_6H_6O_2$, AgBr
(b) $C_6H_6O_2$, AgBr, AgBr, $C_6H_6O_2$
(c) AgBr, $C_6H_6O_2$, AgBr, $C_6H_6O_2$
(d) $C_6H_6O_2$, AgBr, $C_6H_6O_2$, AgBr

154 Consider the reactions,

I. $H_3PO_2(aq) + 4AgNO_3(aq) + 2H_2O(l) \longrightarrow H_3PO_4(aq) + 4Ag(s) + 4HNO_3(aq)$

II. $H_3PO_2(aq) + 2CuSO_4(aq) + 2H_2O(l) \longrightarrow H_3PO_4(aq) + 2Cu(s) + 2H_2SO_4(aq)$

III. $C_6H_5CHO(l) + 2[Ag(NH_3)_2]^+(aq) + 3OH^-(aq) \longrightarrow C_6H_5COO^-(aq) + 2Ag(s) + 4NH_3(aq) + 2H_2O(l)$

IV. $C_6H_5CHO(l) + 2Cu^{2+}(aq) + 5OH^-(aq) \longrightarrow$ No change observed

What inference do you draw about the behaviour of Ag^+ and Cu^{2+} from these reactions?

(a) Ag^+ is weaker oxidising agent than Cu^{2+}
(b) Ag^+ is stronger oxidising agent than Cu^{2+}
(c) Both Ag^+ and Cu^{2+} are stronger oxidising agent
(d) Both Ag^+ and Cu^{2+} are weaker oxidising agent

155 What sort of informations can you draw from the following reaction?

$$(CN)_2(g) + 2OH^-(aq) \longrightarrow CN^-(aq) + CNO^-(aq) + H_2O(l)$$

(a) Decomposition of cyanogen in the cyanide ion (CN^-) and cyanate ion (CNO^-) occurs in basic medium
(b) Cyanogen $(CN)_2$ acts as reducing agent
(c) The reaction is an example of decomposition reaction
(d) All of the above

156 In Ostwald's process for the manufacture of nitric acid, the first step involves the oxidation of ammonia gas by oxygen gas to give nitric oxide gas and steam. What is the maximum weight of nitric oxide that can be obtained starting only with 10.00 g of ammonia and 20.00 g of oxygen?

(a) 10 g (b) 20 g
(c) 16 g (d) 15 g

157 Predict the products of electrolysis of the following. An aqueous solution of $AgNO_3$ with silver electrodes.

(a) Ag from anode dissolves, while Ag^+ ions get reduced and deposited at cathode
(b) Ag from cathode dissolves, while Ag^+ ions get reduced and deposited at anode
(c) Ag^+ from anode dissolves while Ag get reduced and deposited at cathode
(d) Ag^+ from cathode dissolves while Ag^+ ions get oxidised and deposited at anode

158 Arrange the following metals in the order in which they displace each other from the solution of their salts.

Al, Cu, Fe, Mg and Zn.

Given : $E^\circ_{Al^{3+}/Al} = -1.66V; E^\circ_{Cu^{2+}/Cu} = +0.34$ V,
$E^\circ_{Fe^{2+}/Fe} = -0.44V; E^\circ_{Mg^{2+}/Mg} = -2.36V$ and
$E^\circ_{Zn^{2+}/Zn} = -0.76V$

(a) Mg, Zn, Fe, Al, Cu (b) Cu,Zn, Mg, Al, Fe
(c) Mg, Al, Zn, Fe, Cu (d) Al, Mg, Zn, Cu, Fe

159 Given the standard electrode potentials,

$K^+/K = -2.93$ V, $Ag^+/Ag = 0.80$ V,
$Hg^{2+}/Hg = 0.79$ V
$Mg^{2+}/Mg = -2.37$ V, $Cr^{3+}/Cr = -0.74$ V

Arrange these metals in their increasing order of reducing power.

(a) Ag < Cr < Hg < K < Mg
(b) Ag < Cr < Hg < Mg < K
(c) Ag < Hg < Cr < Mg < K
(d) K < Mg < Cr <, < Hg < Ag

NCERT Exemplar

160 Which of the following is not an example of redox reaction?

(a) $CuO + H_2 \longrightarrow Cu + H_2O$
(b) $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$
(c) $2K + F_2 \longrightarrow 2KF$
(d) $BaCl_2 + H_2SO_4 \longrightarrow BaSO_4 + 2HCl$

161 The more positive the value of E^S, the greater is the tendency of the species to get reduced. Using the standard electrode potential of redox couples given below find out which of the following is the strongest oxidising agent.

E° values: $Fe^{3+}/Fe^{2+} = +0.77$; $I_2(s)/I^- = +0.54$;
$Cu^{2+}/Cu = +0.34$; $Ag^+/Ag = +0.80$ V

(a) Fe^{3+} (b) I_2
(c) Cu^{2+} (d) Ag^+

162 E° values of some redox couples are given below. On the basis of these values choose the correct option.

E° values: $Br_2/Br^- = +1.09$; $Ag^+/Ag(s) = +0.80$;
$Cu^{2+}/Cu(s) = +0.34$; $I_2(s)/I^- = +0.54$

(a) Cu will reduce Br^- (b) Cu will reduce Ag
(c) Cu will reduce I^- (d) Cu will reduce Br_2

163 Thiosulphate reacts differently with iodine and bromine in the reactions given below.

$$2S_2O_3^{2-} + I_2 \longrightarrow S_4O_6^{2-} + 2I^-$$
$$S_2O_3^{2-} + 2Br_2 + 5H_2O \longrightarrow 2SO_4^{2-} + 4Br^- + 10H^+$$

Which of the following statements justifies the above dual behaviour of thiosulphate?

(a) Bromine is a stronger oxidant than iodine
(b) Bromine is a weaker oxidant than iodine
(c) Thiosulphate undergoes oxidation by bromine and reduction by iodine in these reactions
(d) Bromine undergoes oxidation and iodine undergoes reduction in these reactions

164 The oxidation number of an element in a compound is evaluated on the basis of certain rules. Which of the following rules is incorrect in this respect?

(a) The oxidation number of hydrogen is always +1
(b) The algebraic sum of all the oxidation numbers in a compound is zero
(c) An element in the free or the uncombined state bears oxidation number zero
(d) In all its compounds, the oxidation number of fluorine is −1

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

(a) NH_2OH (b) NH_4NO_3
(c) N_2H_4 (d) N_3H

166 Which of the following arrangements represent increasing oxidation number of the central atom?

(a) CrO_2^-, ClO_3^-, CrO_4^{2-}, MnO_4^-

(b) ClO_3^-, CrO_4^{2-}, MnO_4^-, CrO_2^-

(c) CrO_2^-, ClO_3^-, MnO_4^-, CrO_4^{2-}

(d) CrO_4^{2-}, MnO_4^-, CrO_2^-, ClO_3^-

167 The largest oxidation number exhibited by an element depends on its outer electronic configuration. With which of the following outer electronic configurations the element will exhibit largest oxidation number?

(a) $3d^1 4s^2$ (b) $3d^3 4s^2$

(c) $3d^5 4s^1$ (d) $3d^5 4s^2$

168 Identify disproportionation reaction,

(a) $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$

(b) $CH_4 + 4Cl_2 \longrightarrow CCl_4 + 4HCl$

(c) $2F_2 + 2OH^- \longrightarrow 2F^- + OF_2 + H_2O$

(d) $2NO_2 + 2OH^- \longrightarrow NO_2^- + NO_3^- + H_2O$

Answers

> Mastering NCERT with MCQs

1 (a)	2 (d)	3 (a)	4 (b)	5 (a)	6 (a)	7 (a)	8 (d)	9 (b)	10 (d)
11 (a)	12 (a)	13 (b)	14 (d)	15 (a)	16 (b)	17 (b)	18 (a)	19 (b)	20 (a)
21 (b)	22 (a)	23 (c)	24 (d)	25 (a)	26 (d)	27 (c)	28 (b)	29 (b)	30 (a)
31 (b)	32 (d)	33 (a)	34 (b)	35 (d)	36 (c)	37 (c)	38 (b)	39 (c)	40 (c)
41 (b)	42 (a)	43 (d)	44 (a)	45 (c)	46 (c)	47 (b)	48 (a)	49 (c)	50 (b)
51 (b)	52 (d)	53 (c)	54 (c)	55 (a)	56 (b)	57 (c)	58 (d)	59 (a)	60 (c)
61 (b)	62 (d)	63 (c)	64 (c)	65 (c)	66 (a)	67 (d)	68 (d)	69 (c)	70 (b)
71 (a)	72 (b)	73 (b)	74 (d)	75 (a)	76 (d)	77 (d)	78 (b)	79 (d)	80 (a)
81 (c)	82 (b)	83 (a)	84 (b)	85 (c)	86 (a)	87 (b)	88 (d)	89 (a)	90 (a)
91 (a)	92 (a)	93 (c)	94 (d)	95 (d)	96 (b)	97 (a)	98 (b)	99 (c)	100 (b)
101 (c)	102 (a)	103 (c)	104 (c)	105 (c)	106 (b)	107 (d)	108 (b)	109 (b)	110 (b)
111 (c)	112 (b)	113 (a)	114 (d)	115 (b)	116 (a)	117 (a)	118 (a)	119 (a)	120 (b)
121 (d)	122 (b)	123 (b)	124 (a)	125 (d)					

> Special Types Questions

126 (c)	127 (d)	128 (a)	129 (b)	130 (a)	131 (b)	132 (b)	133 (c)	134 (d)	135 (c)
136 (a)	137 (a)	138 (a)	139 (a)	140 (b)	141 (c)	142 (b)	143 (d)	144 (b)	145 (b)
146 (a)	147 (d)	148 (c)	149 (a)	150 (c)	151 (b)				

> NCERT & NCERT Exemplar Questions

152 (a)	153 (b)	154 (b)	155 (d)	156 (d)	157 (a)	158 (c)	159 (c)	160 (d)	161 (d)
162 (d)	163 (a)	164 (a)	165 (b)	166 (a)	167 (d)	168 (d)			

Hints & Explanations

3 (*a*) In the given reaction,

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

oxidation of Mg is taking place.

4 (*b*) $CH_2{=\!=}CH_2 + H_2 \longrightarrow H_3C{-}CH_3$

(Addition of hydrogen)

Reduction of ethylene occurs due to the addition of hydrogen.

5 (*a*) In the given reaction,

$$2K_4[Fe(CN)_6](aq) + H_2O_2(aq) \longrightarrow 2K_3[Fe(CN)_6](aq) + 2KOH(aq)$$

Removal of potassium (electropositive element)

So, here oxidation takes place due to removal of one potassium atom.

6 (*a*) In the given reaction, removal of electronegative element, i.e. chlorine from ferric chloride takes place. Hence, it is an example of reduction process.

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

Removal of potassium
(electropositive element)

7 (*a*) $H_2S + Cl_2 \longrightarrow 2HCl + S$

Addition of hydrogen
(reduction)

Removal of hydrogen (oxidation)

Thus, H_2S is oxidised and Cl_2 is reduced.

8 (*d*) Electrochemical processes for the extraction of highly reactive metals and non-metals, manufacturing of chemical compounds like caustic soda, and corrosion of metals fall within the range of redox processes.

9 (*b*) In the given reaction, aluminium is oxidised because oxygen is added to it. Ferrous oxide (Fe_3O_4) is reduced because oxygen has been removed from it.

In other words, Al is changing to Al^{3+} i.e. loss of e^- and Fe^{2+}, Fe^{3+} are converting into Fe, i.e. gain of e^-.

10 (*d*) In the given reaction,

Loss of e^- (Na is oxidised)

$$2\overset{0}{Na}(s) + H_2(g) \longrightarrow 2\overset{+1}{Na}\overset{-1}{H}(s)$$

Gain of e^- (H_2 is reduced)

Thus, Na undergoes oxidation and hydrogen undergoes reduction.

11 (*a*) The half reactions that involve gain of electrons are known as reduction reaction.

12 (*a*) Element which donates electrons is called reducing agent and element which accept electrons is called oxidising agent. Reducing agent reduces other and oxidises itself.

13 (*b*) Valence electrons in Na = 1

Charge on Na = + 1.

Valence electrons in S = 6.

Charge on S = – 2.

So, sodium sulphide (with its charges) is written as $(Na^+)_2S^{2-}$.

Here, –2 charge of S is being neutralised by +1 charge of two Na.

14 (*d*) Loss of $2e^-$

$$2Na + Cl_2 \longrightarrow 2\overset{+}{Na}Cl^-$$

Gain of $2e^-$

15 (*a*) Half-reaction that involves loss of electrons is called oxidation half-reaction. e.g.

$$2Na \longrightarrow 2Na^+ + 2e^-$$

16 (*b*) $4Na + O_2 \longrightarrow 2Na_2O$

Loss of e^- (oxidation)

In the above reaction, Na converts into (Na^+) ion, i.e. Na donates its electron to oxygen atom. So, it behaves as reducing agent.

17 (*b*) $2Na + \overset{0}{S} \longrightarrow (\overset{+}{Na})_2\overset{2-}{S}$

Gain of electrons (reduction)

Thus, sulphur is reduced and sodium is oxidised.

18 (*a*) $H_2S + \overset{0}{I_2} \longrightarrow S + 2H\overset{-1}{I}$

Gain of e^- (I_2 is reduced)

I_2 is reduced and hence acts as oxidising agent.

Thus, the reaction shows oxidising nature of I_2.

19 (*b*) The given reaction is as follows :

$$\overset{0}{Mg}(s) + H_2SO_4(l) \longrightarrow H_2(g) + \overset{+2}{Mg}SO_4(s)$$

Loss of e^- (oxidation takes place)

Thus, oxidation of Mg to $MgSO_4$ takes place.

20 (*a*) $2Na_2S_2O_3 + \overset{0}{I_2} \longrightarrow Na_2S_4O_6 + 2\overset{+1}{Na}\overset{-1}{I}$

Gain of e^- (reduction)

In this reaction, reduction of I_2 to I^- takes place, i.e. gain of e^-. Hence, it acts as oxidising agent.

21 (*b*) $2\overset{+1}{K}\overset{+5}{Cl}\overset{-2}{O_3} \longrightarrow 2\overset{+1}{K}\overset{-1}{Cl} + 3\overset{0}{O_2}$

Loss of e^-
(oxidation)

Gain of e^- (reduction)

Thus, oxygen is oxidised and chlorine is reduced in the above reaction as loss of e^- in oxygen and gain of e^- in Cl takes place.

22 (*a*) When a strip of metallic zinc is placed in an aqueous solution of $Cu(NO_3)_2$, zinc appears as ions (Zn^{2+}).

$$\overset{0}{Zn}(s) + Cu(NO_3)_2(aq) \longrightarrow \overset{+2}{Zn}(NO_3)_2(aq) + Cu(s)$$

23 (*c*) When copper nitrate solution reacts with zinc metal, Cu^{2+} ion is reduced to Cu by gaining electrons from zinc metal.

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

24 (*d*) When rod of metallic zinc is dipped in the solution of copper nitrate, the intensity of blue colour of solution decreases because of the dissolution of zinc rod and finally reddish brown copper metal starts depositing on the zinc rod.

25 (*a*) Another metal (A_1) undergoes oxidation and releases x electron and we know that, for completion of a redox reaction :

Electrons used at cathode (Reduction)
= Electrons released at anode (Oxidation)

Thus, for the reaction

$$A_2^{x+} + xe^- \longrightarrow A_2 \qquad \text{[At cathode]}$$

$$\therefore \quad A_1 \longrightarrow A_1^{x+} + xe^- \qquad \text{[At anode]}$$

Hence, x electrons are released to complete redox reactions.

26 (*d*) Copper being more reactive than silver, can displace silver from its salt solution but opposite is not possible. Thus, equilibrium lies towards right hand side, i.e. in favour of Cu^{2+} and Ag.

27 (*c*) Cobalt (Co) being more reactive than nickel, is oxidised to Co^{2+} and Ni^{2+} ions are reduced to Ni by gaining $2e^-$ from Co.

28 (*b*) Copper is more reactive than silver, it displaces Ag^+ ions from its salt solution which get deposited on the copper rod and Cu^{2+} ions form blue colour.

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

29 (*b*) Zinc releases electrons to copper and copper releases electrons to silver, therefore the electron releasing tendency of the given metals is in the order

$$Zn > Cu > Ag.$$

30 (*a*) Hydrogen is oxidised by loss of electrons

$$2\overset{0}{H_2} + \overset{0}{O_2} \longrightarrow 2\overset{+1}{H_2}\overset{-2}{O}$$

Oxygen is reduced by gain of 2 electrons

In this reaction, hydrogen (H) has transferred electrons to oxygen (O).

31 (*b*) In oxidation number method, it is always assumed that there is a complete transfer of electron from a less electronegative atom to a more electronegative atom.

32 (*d*) Each atom of an element in its free or uncombined state, bears zero oxidation number.

e.g. $\overset{0}{Na}(s), \overset{0}{Ca}(s), \overset{0}{Zn}(s), \overset{0}{Cu}(s), \overset{0}{Cl_2}, \overset{0}{N_2}, \overset{0}{H_2}, \overset{0}{O_2}$ etc.

33 (*a*) For ions composed of only one atom, i.e. monoatomic ions, the oxidation number is equal to the charge on the ion.

e.g. oxidation number of Mg^{2+} ion is + 2,
oxidation number of Fe^{3+} ion is + 3 etc.

34 (*b*) In peroxides (such as H_2O_2, Na_2O_2) each oxygen atom is assigned an oxidation number of −1 and in superoxides (such as KO_2, RbO_2), it is −1/2.

35 (*d*) Electronegativity of fluorine is more than that of oxygen atom, so F gains electron with negative charge. In oxygen difluoride (OF_2) and dioxygen difluoride (O_2F_2), oxygen transfers electron to fluorine atom. Thus,

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

36 (*c*) Algebraic sum of the oxidation number of all the atoms present in a compound must be zero.

e.g. $\overset{+2\,-2\,=\,0}{CaSO_4}$, $\overset{+1\,-1\,=\,0}{NaCl}$, $\overset{+1\,-2}{H_2O}$

$$\downarrow$$

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

37 (*c*) Let the oxidation state of nitrogen in the given compounds be x.

(a) $\overset{x}{N_2}\overset{+1}{H_4}$:

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

(b) $\overset{x}{N}\overset{+1}{H_3}$:

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

(c) $\overset{x}{N_3}\overset{+1}{H}$:

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

(d) $\overset{x}{N}\overset{+1}{H_2}\overset{-2}{O}\overset{+1}{H}$:

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

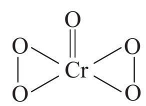
The oxidation state of nitrogen is highest in N_3H.

i.e. in minus, $-\frac{1}{3}$ is greater than −3.

38 (*b*) Let the oxidation number of Cr in $Na_2Cr_2O_7$ be x.

$$\therefore \quad 2(+1) + 2x + 7(-2) = 0$$
$$2 + 2x - 14 = 0$$
$$2x = 12 \Rightarrow x = 6$$

39 (*c*) The structure of CrO_5 is



In $CrO(O_2)_2$, let the oxidation state of Cr be x.

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

Thus, oxidation state of Cr is +6 due to the presence of two peroxide linkages.

40 (*c*) Let the oxidation number of Cl = x.

(i) $ClO_4^- : x + (-2 \times 4) = -1$
$$x - 8 = -1, x = +7$$
(ii) $ClO_3^- : x + (-2 \times 3) = -1$
$$x - 6 = -1, x = +5$$
(iii) $ClO_2^- : x + (-2 \times 2) = -1$
$$x - 4 = -1, \quad x = +3.$$
(iv) $ClO^- : x + (-2) = -1$
$$x - 2 = -1, x = +1$$

Thus, ClO_2^- shows oxidation state of +3.

41 (*b*) $\overset{+5\ -1}{PCl_5}$, $\overset{+5\ -2}{P_2O_5}$, $\overset{+1\ +4\ -2}{H_2PO_3}$

Oxidation no. $\downarrow$ +5, $\downarrow$ +5, $\downarrow$ +4

Thus, the oxidation number of P in PCl_5, P_2O_5 and H_2PO_3 is + 5, + 5 and + 4, respectively.

42 (*a*) The oxidation number of Mn and S in $KMnO_4$ and $Na_2S_2O_3$ respectively are +7 and +2.

It is shown below :

$\overset{+1\ +7\ -2}{KMnO_4}$, $\overset{+1\ +2\ -2}{Na_2S_2O_3}$

Oxidation no. $[+1 + x + (-2)4 = 0$ or $x = +8 - 1 = +7]$ $\quad$ $[2\,H(+1) + 2x + (-2)3 = 0$ or $2x = 6 - 2 = 4]$ $\left[x = \frac{4}{2} = +2\right]$

43 (*d*) Let the oxidation number of carbon in each of the given compounds be x.

(a) $C_6H_{12}O_6 = 6x + 12(+1) + 6(-2) = 0$
$$\Rightarrow \quad x = 0$$
(b) $C_{12}H_{22}O_{11} = 12x + 22(+1) + 11(-2) = 0$
$$\Rightarrow \quad x = 0$$
(c) $CH_4 = x + 4(+1) = 0$
$$\Rightarrow \quad x = -4$$

Thus, oxidation number of carbon is zero in both $C_6H_{12}O_6$ and $C_{12}H_{22}O_{11}$.

44 (*a*) The highest value of oxidation number exhibited by an atom of an element generally increases across a period in the periodic table.

In the third period, the highest value of oxidation number changes from 1 to 7.

45 (*c*) For the given compounds,

(I) $\overset{+1\ +6\ -2}{K_2MnO_4}$, (II) $\overset{+5\ -2}{Mn_2O_5}$, (III) $\overset{+7\ -2}{MnO_4^-}$, (IV) $\overset{+2}{MnCl_2}$

Oxidation no. $\downarrow$ +6, $\downarrow$ +5, $\downarrow$ +7, $\downarrow$ +2

Decreasing order of oxidation number is

$$III > I > II > IV$$

46 (c) Let the oxidation state of nitrogen in each of the given N-compounds be x.

(i) $HNO_3 : +1 + x + 3(-2) = 0$
$$x = +5$$
$\therefore$ Oxidation state of N in HNO_3 is +5.

(ii) $NO : x + 1(-2) = 0$
$$x = +2$$
$\therefore$ Oxidation state of N in NO is +2.

(iii) $NH_4Cl : x + 4(+1) + 1(-1) = 0$
$$x = -3$$
$\therefore$ Oxidation state of N in NH_4Cl is -3.

(iv) $N_2 : x = 0$ [$\because N_2$ is present in elemental state]
$\therefore$ Oxidation state of N in N_2 molecule is 0.

Thus, the correct decreasing order of oxidation states of given N- compounds will be

$$\overset{+5}{H\,N}O_3 > \overset{+2}{N}O > \overset{0}{N_2} > \overset{-3}{N}H_4Cl$$

47 (*b*) Loss of e^- (oxidation)

$$H_2O + \overset{0}{Br_2} \longrightarrow \overset{+}{H}\,\overset{-}{Br} + H\overset{-}{O}\overset{+}{Br}$$

Gain of e^- (reduction)

Here, oxidation number of bromine increases as well as decreases, i.e. bromine is oxidised as well as reduced.

48 (*a*) $O{=}\overset{+2}{C}{=}\overset{0}{C}{=}\overset{+2}{C}{=}O$

Carbon suboxide

In C_3O_2 , two C-atoms linked with oxygen atoms are present in +2 oxidation state and central carbon has zero oxidation state.

So, the average oxidation state of carbon is

$$= \frac{+2 + 0 + 2}{3} = +\frac{4}{3}.$$

49 (*c*) $\underset{+6}{H_2SO_5}$, $\underset{+6}{Cr_2O_7^{2-}}$, $\underset{+5}{NO_3^-}$

Oxidation no. +6, +6, +5

$$2 + x + (-2\times 3) + (-1\times 2) = 0, \quad 2x + (-2\times 7) = -2, \quad x + (-2\times 3) = -1$$
$$x = +6 \qquad x = +6 \qquad x = +5$$

Thus, the oxidation number of S, Cr and N in H_2SO_5, $Cr_2O_7^{2-}$ and NO_3^-, respectively are +6, + 6 and +5.

Note *In the structure of* H_2SO_5, *two oxygen atoms form peroxide linkage, so oxidation number of these O-atoms is* −1.

50 (*b*) NH_3, HNO_3, NaN_3, Mg_3N_2

Oxidation no. −3, +5, −1/3, −3

Hence, in three molecules (NH_3, NaN_3 and Mg_3N_2) nitrogen have negative oxidation state and in one molecule (HNO_3) nitrogen has positive oxidation state.

51 (*b*) We know that, the algebraic sum of the oxidation states is always zero in neutral compound.

Oxidation states of $X = +2$

$$Y = +5$$
$$Z = -2$$

So, the algebraic sum of total *X*, *Y* and *Z* should be equal to zero which is found in XY_2Z_6.

$$XY_2Z_6 = +2 + (5\times 2) + (-2\times 6)$$
$$= +2 + 10 - 12 = 0$$

52 (*d*) Given, electronic configuration of *X*

$$= 1s^2 2s^2 2p^3$$

∴ The valency of *X* will be 3.

The valency of Mg is +2.

∴ Magnesium reacts with element *X* to form an ionic compound with formula Mg_3X_2.

53 (*c*) Let the oxidation number of Br is *x*.

$$Br_3O_8$$
$$3x + (-2)\times 8 = 0$$
$$3x + (-16) = 0$$
$$\Rightarrow \quad x = \frac{16}{3}$$

Thus, the average oxidation number of Br in Br_3O_8 is $\frac{16}{3}$.

54 (*c*) The oxidation state of a metal in a compound is represented according to the notation which is known as Stock notation.

55 (*a*) By using Stock-notation, the given compounds can be represented as,

$$FeO \longrightarrow Fe(II)O$$
$$Fe_2O_3 \longrightarrow Fe_2(III)O_3$$
$$CuO \longrightarrow Cu(II)O$$
$$MnO_2 \longrightarrow Mn(IV)O_2$$

where, II, III, II, IV represent oxidation states of metals, Fe (in FeO), Fe (in Fe_2O_3), Cu (in CuO) and Mn (in MnO_2) respectively.

56 (*b*) An increase in the oxidation number of the element in the given substance is known as oxidation.

e.g. $Fe^{2+} \longrightarrow Fe^{3+} + e^-$

57 (*c*) A reagent which lowers the oxidation number of an element in a given substance is known as reducing agent or reductant.

$$Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$$

Here, Zn is called as reducing agent.

59 (*a*)

Loss of e^- (oxidation)

$$2\overset{+1}{Cu_2}\overset{-2}{O}(s) + \overset{+1}{Cu_2}\overset{-2}{S}(s) \longrightarrow 6\overset{0}{Cu}(s) + \overset{+4}{S}\overset{-2}{O_2}(g)$$

Gain of e^- (reduction)

Thus, copper is reduced and sulphur is oxidised.

60 (*c*)

Oxidation (Loss of e^-)

$$2\underset{+2}{Hg}Cl_2(aq) + \underset{+2}{Sn}Cl_2(aq) \longrightarrow \underset{+1}{Hg_2}Cl_2(s) + \underset{+4}{Sn}Cl_4(aq)$$

Reduction Gain of e^-

The given reaction is an example of redox reaction because oxidation of stannous chloride to stannic chloride and reduction of mercuric chloride to mercurous chloride take place simultaneously.

61 (*b*) In the reaction,

$$2\underset{+1}{Cu_2}O + \underset{+1}{Cu_2}S \longrightarrow 6\underset{0}{Cu} + SO_2\uparrow$$

sulphur of Cu_2S acts as reductant because it decreases the oxidation number of copper from + 1 to 0.

62 (*d*) The change in oxidation number of underlined substance in the given reactions are as follows :

(a) $3CaO + \overset{0}{N_2} \longrightarrow Ca_3\overset{-3}{N_2}$

Decrease in oxidation number (Gain of e^-)
(N_2 is reduced)

(b) $2NaI + \overset{0}{Br_2} \longrightarrow 2\overset{+1}{Na}\overset{-1}{Br} + I_2$

Decrease in oxidation number (Gain of e^-)
(Br_2 is reduced)

(c) $\overset{+2}{Zn}\overset{-2}{O} + H_2 \longrightarrow \overset{0}{Zn} + H_2O$

Decrease in oxidation number (Gain of e^-)
(ZnO is reduced)

(d) $\overset{+2}{CO} + Cl_2 \longrightarrow \overset{+4}{C}\overset{-2}{O}\overset{-1}{Cl_2}$

Increase in oxidation number (Loss of e^-)
(CO is oxidised)

Thus, , CO is oxidised among the given reactions.

63 (*c*) $\overset{+1}{K}\overset{+7}{Mn}\overset{-2}{O_4} + H_2SO_4\,\text{(Acidified solution)} \longrightarrow \overset{2+}{Mn}SO_4$

Oxidation number of Mn is changed from +7 to +2.

64 (*c*) $C_6H_6 + H\overset{+5}{N}O_3 \longrightarrow C_6H_5 - \overset{+3}{N}O_2 + H_2O$

Reduction (Gain of e^-)

In this reaction, HNO_3 behaves as an oxidising agent while in rest of the reactions such as,

(i) $NaOH + HNO_3 \longrightarrow NaNO_3 + H_2O$

(ii) $Ca(OH)_2 + 2HNO_3 \longrightarrow Ca(NO_3)_2 + 2H_2O$

(iii) $NaCl + HNO_3 \longrightarrow HCl + NaNO_3$

HNO_3 neither behaves as oxidising agent nor as reducing agent.

65 (*c*) $\underset{+6}{Cr_2}O_7^{2-} + \underset{+1}{\overset{+1}{H_2}}\overset{-2}{O} \longrightarrow 2\underset{+6}{Cr}O_4^{2-} + 2\underset{+1}{H^+}$

Oxidation states of chromium in $Cr_2O_7^{2-}$ and CrO_4^{2-} is +6

Similarly, oxidation states of hydrogen in H_2O and H^+ is +1. Thus, oxidation number of chromium has neither increased nor decreased.

66 (*a*) All combustion reactions, which make use of elemental dioxygen are redox reactions.

e.g. $C(s) + O_2(g) \longrightarrow CO_2(g)$

67 (*d*) The reaction in which an element or a compound reacts with oxygen is called combustion reaction.

$$C(s) + O_2(g) \xrightarrow{\Delta} CO_2(g)$$

$$CH_4(g) + 2O_2(g) \xrightarrow{\Delta} CO_2(g) + 2H_2O(l)$$

68 (*d*) A combination reaction may be denoted in the following manner :

$$A + B \longrightarrow C$$

Either A and B or both A and B must be in the elemental form for such a reaction to be redox reaction.

69 (*c*) When a compound dissociates into two or more components atleast one of which must be in the elemental state, it refers to as a decomposition reaction.

70 (*b*) $2\overset{+1}{H_2}\overset{-2}{O}(l) \xrightarrow{\Delta} 2\overset{0}{H_2}(g) + \overset{0}{O_2}(g)$

The above reaction involves decomposition of H_2O molecule into H_2 and O_2.

71 (*a*) The reaction in which oxidation and reduction occur simultaneously are termed as redox reaction.

$$(\overset{+4}{X})eF_4 + \overset{+1}{O_2}(F_2) \longrightarrow \overset{+6}{X}eF_6 + \overset{0}{O_2}$$

Since, Xe undergoes oxidation while O undergoes reduction. So, it is an example of redox reaction.

72 (*b*) $\overset{+5}{P}\overset{-1}{Cl_5} \xrightarrow{\Delta} \overset{+3}{P}Cl_3 + \overset{0}{Cl_2}$

Gain of e^-

(reduction)

In the above reaction, same compound, i.e. PCl_5 is oxidised and reduced, so it is an example of intramolecular redox change.

73 (*b*) $2\overset{+1}{K}\overset{+5}{Cl}\overset{-2}{O_3} \longrightarrow 2\overset{+1}{K}\overset{-1}{Cl} + 3\overset{0}{O_2}$

(a) The oxidation number of K does not change, thus K undergoes neither reduction nor oxidation.

(b) The oxidation number of chlorine decreases from +5 in $KClO_3$ to –1 in KCl, hence Cl undergoes reduction.

(c) Since, oxidation number of oxygen increases from –2 in $KClO_3$ to 0 in O_2, so oxygen is oxidised.

(d) This statement is not correct because Cl is undergoing reduction and O is undergoing oxidation.

Therefore, statement (b) is true about the given reaction.

74 (*d*) Compounds, e.g. Mn_2O_7 when dissociats into O_2 and MnO_2, oxidation and reduction both take place.

i.e. $2\overset{+5}{M}n_2\underset{-2}{O_7} \longrightarrow 4\overset{+4}{Mn}O_2 + 3\underset{0}{O_2}$

It is an example of intramolecular redox reaction, while reaction, such as

Gain of e^-

$\underset{+2}{Sn}Cl_2 + 2\overset{+3}{Fe}Cl_3 \longrightarrow \underset{+4}{Sn}Cl_4 + 2\overset{+2}{Fe}Cl_2$

Loss of e^-

is an example of intermolecular redox reaction.

75 (*a*) $\overset{2+}{Cu}\overset{+6}{S}\overset{-2}{O_4} + \overset{0}{Zn} \longrightarrow \overset{0}{Cu} + \overset{2+}{Zn}\overset{+6}{S}\overset{-2}{O_4}$

In this reaction, copper (Cu^{2+}) is displaced by zinc metal, so this reaction is called metal displacement reaction.

76 (*d*) Reactions in which the same substance is oxidised as well as reduced are called disproportionation reaction.

Loss of e^-

$\overset{-4}{C}\overset{+1}{H_4} + 2\overset{0}{O_2} \longrightarrow \overset{+4}{C}\overset{-2}{O_2} + 2\overset{+1}{H_2}\overset{-2}{O}$

Loss of e^-

$\overset{-4}{C}\overset{+1}{H_4} + 4\overset{0}{Cl_2} \longrightarrow \overset{+4}{C}\overset{-1}{Cl_4} + 4\overset{+1}{H}\overset{-1}{Cl}$

Gain of e^-

$2\overset{0}{F_2} + 2\overset{-2}{O}\overset{+1}{H^-} \longrightarrow 2\overset{-1}{F^-} + \overset{+2}{O}\overset{-1}{F_2} + \overset{+1}{H_2}\overset{-2}{O}$

Loss of e^-

$2\overset{+4}{N}\overset{-2}{O_2} + 2\overset{-2}{O}\overset{+1}{H^-} \longrightarrow \overset{+3}{N}\overset{-2}{O_2^-} + \overset{+5}{N}\overset{-2}{O_3^-} + \overset{+1}{H_2}\overset{-2}{O}$

Gain of e^-

Thus, in the reaction of NO_2 and OH^-, N is both oxidised as well as reduced. Hence, the oxidation number increases from +4 in NO_2 to +5 in NO_3^- and decreases from +4 in NO_2 to +3 in NO_2^-.

77 (*d*) The reactions in which the same species is simultaneously oxidised and reduced are called disporportionation reactions.

Let us, consider the given reaction one by one :

(i) $2Cu^+ \longrightarrow Cu^{2+} + Cu^0$

The above reaction is a disproportionation reaction as Cu(+1) is oxidised to Cu(+2) and reduced to Cu(0).

(ii) $3\overset{+6}{Mn}O_4^{2-} + 4H^+ \longrightarrow 2\overset{+7}{Mn}O_4^- + \overset{+4}{Mn}O_2 + 2H_2O$

The above reaction is a disproportionation reaction as Mn(+ 6) is oxidised to $MnO_4^-(Mn^{+7})$ and reduced to $MnO_2(\overset{+4}{Mn})$.

(iii) $2K\overset{+7}{Mn}O_4 \xrightarrow{\Delta} K_2\overset{+6}{Mn}O_4 + \overset{+4}{Mn}O_2 + O_2$

The above reaction is not a disproportionation reaction as Mn (+7) is only reduced to $K_2MnO_4(Mn^{+6})$ and $MnO_2(Mn^{+4})$.

(iv) $2\overset{+7}{Mn}O_4^- + 3Mn^{2+} + 2H_2O \longrightarrow 5\overset{+4}{Mn}O_2 + 4H^+$

The above reaction is not a disproportionation reaction as Mn(+7) is only reduced to MnO_2(+ 4).

Hence, option (d) is correct.

78 (*b*) All alkali metals and some alkaline earth metals (like Ca, Sr and Ba) which are very good reductants, will displace hydrogen from cold water.

$$2Na + 2HOH \longrightarrow 2NaOH + H_2$$

79 (*d*) The given reactions are :

$$\text{I. } \overset{0}{Zn} + 2\overset{+1\,-1}{HCl} \longrightarrow \overset{+2\,-1}{ZnCl_2} + \overset{0}{H_2}$$

$$\text{II. } \overset{0}{Fe} + 2\overset{+1\,-1}{HCl} \longrightarrow \overset{+2\,-1}{FeCl_2} + \overset{0}{H_2}$$

In both reactions I and II, hydrogen from acid is displaced by zinc and iron metals, respectively.

80 (*a*) Less active metals like magnesium and iron reacts with steam to produce dihydrogen gas.

$$Mg(s) + 2H_2O(l) \xrightarrow{\Delta} Mg(OH)_2(s) + H_2(g)$$

$$2Fe(s) + 3H_2O(l) \xrightarrow{\Delta} Fe_2O_3(s) + 3H_2(g)$$

81 (*c*) Cadmium and tin being less reactive do not react with steam.

82 (*b*) Silver (Ag) and gold (Au) are less reactive metals or noble metals and hence, do not react with HCl.

83 (*a*) The oxidising power decreases from fluorine to iodine in group-17 of the periodic table. This implies that fluorine is so reactive that it can replace chloride, bromide and iodide ions in solution. Thus, the order will be $F_2 < Cl_2 < Br_2 < I_2$

84 (*b*) Phosphorus, sulphur and chlorine disproportionate in alkaline medium as shown below :

Oxidation

$$\overset{0}{P_4}(s) + 3O\bar{H}(aq) + 3H_2O(l) \longrightarrow \overset{-3}{P}H_3(g) + 3H_2\overset{+1}{P}O_2^-$$

Reduction

Reduction

$$\overset{0}{S_8} + 12O\bar{H}(aq) \longrightarrow 4\overset{-2}{S^{2-}}(aq) + 2\overset{+2}{S_2}O_3^{2-}(aq) + 6H_2O$$

Oxidation

Oxidation

$$\overset{0}{Cl_2}(g) + 2O\bar{H}(aq) \longrightarrow \overset{+1}{Cl}O^-(aq) + \overset{-1}{Cl^-}(aq) + H_2O(l)$$

Reduction

85 (*c*) $\overset{0}{Cl_2}(g) + 2OH^-(aq) \longrightarrow \overset{+1}{Cl}O^-(aq) + Cl^-(aq) + H_2O(l)$,

is the reaction which describes the formation of household bleaching agents. The hypochlorite ion (ClO^-) formed in the reaction oxidises the colour-bearing stains of the substances to colourless compounds.

86 (*a*) Fluorine does not show disproportionation tendency because its oxidation state is always negative.

Oxidation (Loss of e^-)

87 (*b*) $\overset{0}{P_4} + 3NaOH + 3H_2O \longrightarrow \overset{-3}{P}H_3 + 3\overset{+1}{Na}H_2\overset{+1}{P}\overset{-2}{O_2}$

Reduction (Gain of e^-)

The oxidation number of phosphorus in PH_3 and NaH_2PO_2 is –3 and +1, respectively, i.e. here phosphorus is oxidised as well reduced. Therefore, this reaction is known as disproportionation reaction.

88 (*d*) In disproportionation reactions, same element undergoes oxidation as well as reduction.

Among the given reaction, reaction (d) is a disproportionation reaction.

Reduction

$$2\overset{+1}{Cu}Br \longrightarrow \overset{+2}{Cu}Br_2 + \overset{0}{Cu}$$

Oxidation

Here, CuBr get oxidised to $CuBr_2$ and also it get reduced to Cu.

89 (*a*) ClO_4^- does not show disproportionation reaction because oxoanion of chlorine is present in its highest oxidation state, i.e. +7, so it does not further increase its oxidation number.

90 (*a*) Fluorine does not show a disproportionation reaction.

$$2\overset{0}{F_2} + 2OH^- \longrightarrow 2\overset{-1}{F^{-1}} + \overset{+2-1}{OF_2} + H_2O$$

i.e. Here, only reduction is taking place no oxidation.

91 (*a*) If a reaction is carried out in acidic medium, H^+ ions are used to balance the equation. If it is carried out in basic medium, OH^- ions are used.

92 (*a*) $\overset{}{Fe^{2+}}(aq) + \overset{+6\ -2}{Cr_2O_7^{2-}}(aq) \longrightarrow Fe^{3+}(aq) + Cr^{3+}(aq)$

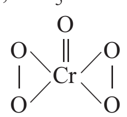
Loss of electron (oxidation)

In the above reaction, Fe^{2+} is getting oxidise to Fe^{3+}.

93 (*c*) When H_2O_2 is added to an acidified solution of a dichromate, $Cr_2O_7^{2-}$, a deep blue coloured complex, chromic peroxide, CrO_5 [or $CrO(O_2)_2$] is formed.

$$Cr_2O_7^{2-} + 2H^+ + 4H_2O_2 \longrightarrow \underset{\text{Chromic peroxide}}{2CrO(O_2)_2} + 5H_2O$$

Chromic peroxide, CrO_5 has the following structure:



Oxidation state of Cr is + 6 due to the presence of two peroxide linkages, which can be calculated as:

In $CrO(O_2)_2$, let the oxidation state of Cr be x.

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

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

94 (*d*) $\overset{6+}{Cr_2O_7^{2-}} \longrightarrow Cr^{3+}$

Gain of e^- (reduction)

Reduction takes place due to decrease in oxidation number, so chromium is being reduced.

95 (*d*) Electrons are added to one side of the half-reaction to balance the charges.

96 (*b*) $\overset{+5}{Bi}O_3^- + 6H^+ + xe^- \longrightarrow Bi^{3+} + 3H_2O$

The half cell reaction can be written as :

$$Bi^{5+} \xrightarrow{+2e^-} Bi^{3+}$$

$$\overset{+5}{Bi}\overset{-2}{O_3^-} + 6H^+ + 2e^- \longrightarrow Bi^{3+} + 3H_2O$$

Thus, the value of $x = 2$ in the above ionic equation.

Also, by another way

$$BO_3^- + 6H^+ + xe^- \longrightarrow B^{3+} + 3H_2O$$

Total charge on LHS = Total charge on RHS

$$-1 + 6 + x(-1) = +3$$

$$-x = +3 - 5$$

$$-x = -2 \Rightarrow x = 2$$

97 (*a*) For the given reaction, oxidation and reduction half-reactions are :

Oxidation half reaction : $6I^- \longrightarrow 3I_2 + 6e^-$

Reduction half reaction :

$$2MnO_4^- + 4H_2O + 6e^- \longrightarrow 2MnO_2 + 8\bar{O}H$$

On adding the above two half-reactions, we get ionic equation,

$$6I^- + 2MnO_4^- + 4H_2O \longrightarrow 3I_2 + 2MnO_2 + 8\bar{O}H$$

Thus, the correct representation of given molecular equation in ionic form is given in option (a)

98 (*b*) The balanced equation is,

$$Cu + \underset{x}{4}HNO_3 \longrightarrow Cu(NO_3)_2 + \underset{y}{2}NO_2 + \underset{z}{2}H_2O$$

$\therefore x = 4,\ y = 2,\ z = 2$

99 (*c*) $\underset{1}{a}\ \overset{+1\ +6\ -2}{K_2Cr_2O_7} + \underset{4}{b}\ \overset{+1}{K}Cl + \underset{6}{c}\ H_2SO_4$

$$\longrightarrow \underset{2}{x}\ \overset{+6\ -2\ -2}{CrO_2Cl_2} + \underset{6}{y}\overset{+1\ +1\ -2}{KHSO_4} + \underset{3}{z}H_2O$$

The balanced equation is,

$$K_2Cr_2O_7 + 4KCl + 6H_2SO_4 \longrightarrow 2CrO_2Cl_2 + 6KHSO_4 + 3H_2O$$

$a = 1, b = 4, c = 6, x = 2, y = 6, z = 3$

100 (*b*) The given redox reaction is

$$MnO_4^- + C_2O_4^{2-} + H^+ \longrightarrow Mn^{2+} + CO_2 + H_2O$$

The reaction can be balanced by considering the following steps;

***Step* I** Balance the atoms except H and O.

$$MnO_4^- + C_2O_4^{2-} + H^+ \longrightarrow Mn^{2+} + 2CO_2 + H_2O$$

***Step* II** Write the oxidation number of each atom

Oxidation ($-2e^-$)

$$\overset{+7}{Mn}O_4^- + \overset{+3}{C_2}O_4^{2-} + H^+ \longrightarrow Mn^{2+} + 2\overset{+4}{C}O_2 + H_2O$$

Reduction ($+5e^-$)

***Step* III** Cross multiply by change in oxidation number

$$\overset{+7}{Mn}O_4^- \longrightarrow Mn^{2+};\ 5e^- \text{ gain}$$

$$\overset{+3}{C_2}O_4^{2-} \longrightarrow 2\overset{+4}{C}O_2;\ 2x\ 1e^- \text{ loss}$$

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

***Step* IV** Balance oxygen by adding H_2O on deficient site.

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

***Step* V** Balance hydrogen

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

$\therefore$ The coefficients of the reactants, MnO_4^-, $C_2O_4^{2-}$ and H^+ are 2, 5 and 16, respectively.

101 (*c*) The given redox reaction is,

$$Fe^{2+} + Cr_2O_7^{2-} + H^+ \longrightarrow Fe^{3+} + 2Cr^{3+} + H_2O$$

The reaction can be balanced by considering the following steps :

***Step* I** Balance the atoms except H and O.

$$Fe^{2+} + Cr_2O_7^{2-} + H^+ \longrightarrow Fe^{3+} + 2Cr^{3+} + H_2O$$

***Step* II** Write the oxidation number of each atom

Oxidation $(-1e^-)$

$$Fe^{2+} + \overset{+6}{Cr_2}O_7^{2-} + H^+ \longrightarrow Fe^{3+} + 2Cr^{3+} + H_2O$$

Reduction $(+6e^-)$

***Step* III** Cross multiply by change in oxidation number.

$$Fe^{2+} \longrightarrow Fe^{3+};\ 1e^- \text{ loss}$$

$$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+};\ 6e^- \text{ gain}$$

$$6Fe^{2+} + Cr_2O_7^{2-} \longrightarrow 6Fe^{3+} + 2Cr^{3+}$$

***Step* IV** Balance oxygen by adding H_2O on deficient site.

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

***Step* V** Balance hydrogen

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

$\therefore$ The coefficients x, y and z are 6, 1 and 14 respectively.

102 (*a*) The equation when BrO_3^- changes into Br_2 in an acidic medium is as follows :

$$2BrO_3^- + 12H^+ \longrightarrow Br_2 + 6H_2O$$

Add $10e^-$ in LHS to balance the charge on both sides

$$2BrO_3^- + 10e^+ + 12H^+ \longrightarrow Br_2 + 6H_2O.$$

$\therefore$ 10 electrons should be present on the left side of balanced equation.

103 (*c*) $K_2Cr_2O_7 + KI \text{ (excess)} \longrightarrow Cr^{3+} + H_2O + nI_2$

$$\overset{6+}{Cr} \xrightarrow{+3e^-} \overset{3+}{Cr}$$

Cr^{6+} accepts $3e^-$, so mole of I_2 liberated = 3.

Thus, the value of n is 3.

104 (*c*) $\overset{+7}{Mn}O_4^-(aq) + I^-(aq) \longrightarrow \overset{+4}{Mn}O_2(s) + \overset{0}{I_2}(s)$

This reaction is represented in basic medium because in basic medium $\overset{+7}{Mn}O_4^-$ is reduced to $\overset{+4}{Mn}O_2$ (i.e. $\overset{+7}{Mn}$ to $\overset{+4}{Mn}$), while in acidic medium, MnO_4^- is reduced from Mn^{7+} to Mn^{2+}.

105 (*c*) When a mixture of potassium chlorate, oxalic acid and sulphuric acid is heated, the following reaction occurs :

$$\overset{+1\,+5\,-2}{KClO_3} + \overset{+1\,+3\,-2}{H_2C_2O_4} + \overset{+1\,+6\,-2}{H_2SO_4} \xrightarrow{\Delta} \overset{+1\,+6\,-2}{K_2SO_4} + \overset{+1\,-1}{KCl} + \overset{+4\,-2}{CO_2} + \overset{+1\,-2}{H_2O}$$

Thus, Cl is the element which undergoes maximum change in the oxidation state from +5 to −1.

106 (*b*) $\overset{+7}{Mn}O_4^-(aq) + I^-(aq) \longrightarrow \overset{+4}{Mn}O_2(s) + \overset{0}{I_2}(s)$

For the above given reaction, the oxidation and reduction half reactions are as follows :

Oxidation half reaction: $I^-(aq) \longrightarrow I_2(s)$

Reduction half reaction: $MnO_4^-(aq) \longrightarrow MnO_2(s)$

So, the oxidation half reaction is given in option (b).

107 (*d*) In redox reaction, the strength of oxidant/reductant is determined by titration method using a redox sensitive indicator.

108 (*b*) Titration method in acid-base system is used for finding out the strength of one solution against the other solution by using pH sensitive indicator.

109 (*b*) In a solution of $KMnO_4$, MnO_4^- acts as a self indicator because in redox titrations, these reagents which are itself has intense colour, act as self indicator.

110 (*b*) $Cr_2O_7^{2-}$ is not a self indicator, but it oxidises the indicator and act as oxidising agent.

111 (*c*) The reaction of Cu^{2+} ions with iodide ions gives an intense blue colour when starch is added. When $S_2O_3^-$ is added the blue colour of solution gets disappeared.

112 (*b*) If zinc rod is dipped in copper sulphate solution, then due to transfer of electrons from zinc to copper ions, heat is evolved.

113 (*a*) The couple having oxidised and reduced forms of a substances taking part in an oxidation or reduction half-reaction is called a redox couple.

114 (*d*) $Zn^{2+} + 2e^- \longrightarrow Zn$

Redox couple = Zn^{2+}/Zn

$$Cu^{2+} + 2e^- \longrightarrow Cu$$

Redox couple = Cu^{2+}/Cu

115 (*b*) In Daniell cell, electrons flow from anode to cathode and current flows from cathode to anode.

116 (*a*) In Daniell cell, electrons are produced at the anode due to oxidation of Zn and at cathode electrons are absorbed due to reduction of copper (Cu^{2+}).

i.e. (i) $Zn \longrightarrow Zn^{2+} + 2e^-$ (at anode)

(ii) $Cu^{2+} + 2e^- \longrightarrow Cu$ (at cathode)

117 (*a*) A salt bridge is a U-tube containing a solution of potassium chloride or ammonium nitrate solidified by boiling with agar-agar and later cooling to a jelly like substance.

118 (*a*) The rods of transition metals such as copper and zinc where potential difference is generated are termed as electrodes.

119 (a) In Daniell cell, negative terminal (p) is anode whereas positive terminal (q) is cathode. Here 'r' is a salt bridge. As electrons are getting transferred from Zn to Cu^{2+}, so 's' will be electron flow. The flow of electricity is always in opposite direction to that of electron flow, therefore 't' is the current flow.

120 (b) A negative $E°$ means that the redox couple is a stronger reducing agent than H^+/H_2 couple. A positive $E°$ means that the redox couple is a weaker reducing agent than the H^+/H_2 couple.

121 (d) The reducing agent oxidises itself by undergoing oxidation through the loss of electrons. Thus, reducing agent reduces other molecules by supplying electrons to them.

(1) $H_2O_2^{-1} + 2H^+ + 2e^- \longrightarrow 2H_2\overset{-2}{O}$

In this reaction, H_2O_2 undergoes reduction as O shows increase in oxidation number.

(2) $H_2\overset{-1}{O}_2 - 2e^- \longrightarrow \overset{0}{O}_2 + 2H^+$

In this reaction, H_2O_2 undergoes oxidation as it shows decrease in oxidation number. Thus, H_2O_2 acts as a reducing agent.

(3) $H_2\overset{-1}{O}_2 + 2e^- \longrightarrow 2\overset{-2}{O}H^-$

In this reaction, H_2O_2 undergoes reduction.

(4) $H_2\overset{-1}{O}_2 + 2OH^- - 2e^- \longrightarrow \overset{0}{O}_2 + 2H_2O$

In this reaction, H_2O_2 undergoes reduction. Thus, acts as a reducing agent.

122 (a) Given,

$Co^{3+} + e^- \longrightarrow Co^{2+}; E° = +1.81\ V$

$Pb^{4+} + 2e^- \longrightarrow Pb^{2+}; E° = +1.67\ V$

$Ce^{4+} + e^- \longrightarrow Ce^{3+}; E° = +1.61\ V$

$Bi^{3+} + 3e^- \longrightarrow Bi; E° = +0.20\ V$

Oxidising power of the species will increase in the order of $Bi^{3+} < Ce^{4+} < Pb^{4+} < Co^{3+}$.

Higher the emf value, stronger the oxidising power. The maximum value of emf is possessed by Co^{3+}. Hence, it has maximum oxidising power. Whereas Bi^{3+} possess the lowest emf value. Hence, it has minimum oxidising power.

123 (b) More negative is the value of standard reduction potential, higher is the reduction power.

i.e. Reducing power $\propto \dfrac{1}{\text{standard reduction potential}}$

Thus, the correct decreasing order of reducing power of the metal is :

$$\underset{(E^o_{K^+/K} = -2.93V)}{K} > \underset{(E^o_{Al^{3+}/Al} = -1.66V)}{Al} > \underset{(E^o_{Cr^{3+}/Cr} = -0.74V)}{Cr} > \underset{(E^o_{Ag^+/Ag} = 0.80V)}{Ag}$$

124 (a) The substances which have lower reduction potentials are strong reducing agents. Therefore, $Cr(E^\circ_{Cr^{3+}/Cr} = -0.74\ V)$ is the strongest reducing agent among all the other given options.

125 (d) For a reaction to be feasible, E°_{cell} must be positive and $E^\circ_{cell} = E^\circ_{oxi} + E^\circ_{red}$

(a) Oxidation half-reaction

$2I^-(aq) \longrightarrow I_2(s) + 2e^-; \quad E° = -0.54\ V$

Reduction half-reaction

$2Fe^{3+}(aq) + 2e^- \longrightarrow 2Fe^{2+}(aq); \ E° = +0.77\ V$

Overall reaction,

$2Fe^{3+}(aq) + 2I^-(aq) \longrightarrow 2Fe^{2+}(aq) + I_2(s);$
$E° = +0.23\ V$

Positive emf indicates that the reaction is feasible.

(b) Oxidation half reaction,

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

Reduction half reaction,

$2Ag^+(aq) + 2e^- \longrightarrow 2Ag(s); \ E° = +0.80\ V$

Overall reaction,

$Cu(s) + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag;$
$E° = +0.46\ V$

Positive emf indicates that the reaction is feasible.

(c) Oxidation half-reaction,

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

Reduction half-reaction,

$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)] \times 2; \ E° = +0.77V$

Overall reaction,

$Cu(s) + 2Fe^{3+}(aq) \longrightarrow Cu^{2+}(aq) + 2Fe^{2+}(aq);$
$E° = +0.43V$

Positive emf indicates that the reaction is feasible.

(d) Oxidation half-reaction,

$Ag(s) \longrightarrow Ag^+(aq) + e^-; E° = -0.80V$

Reduction half-reaction,

$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq); \ E° = +0.77V$

Overall reaction,

$Ag(s) + Fe^{3+}(aq) \longrightarrow Ag^+(aq) + Fe^{2+}(aq);$
$E° = -0.03V$

Negative emf indicates that the reaction is not feasible.

Therefore, reaction is not feasible if the electrode potential are $Ag(s)$ and $Fe^{3+}(aq)$.

126 (c) Statement (c) is incorrect.
It's correct form is as follows :

Ne is an inert gas, so it exhibits neither negative nor positive oxidation state.
Rest other statements are correct.

127 (*d*) Statement (d) is correct, while the other statements are incorrect. Corrected form are as follows :

$$\overset{0}{Zn} + 2\overset{+1}{H}\overset{-1}{Cl} \xrightarrow{} \overset{+2}{Zn}\overset{-1}{Cl_2} + \overset{0}{H_2}$$

(Zn loses of $2e^-$, Oxidation)

The oxidation number of Zn increases from 0 in Zn to +2 therefore, Zn acts as a reductant. Oxidation number of H decreases from +1 to 0, so it acts as an oxidant.

128 (*a*) Statement (a) is corrects while the other statements are incorrect. Corrected form are as follows :

In the given reaction,

$$3\overset{-2}{N_2}H_4 + 2\overset{+5}{Br}O_3^- \longrightarrow 3\overset{0}{N_2} + 2\overset{-1}{Br^-} + 6H_2O$$

Oxidation number of N changes from −2 to 0, it is oxidised and acts as a reducing agent. Oxidation number of Br changes from +5 to −1, it is reduced and acts as an oxidising agent.

129 (*b*) Statement (b) is incorrect.

It's correct form is as follows :

Oxidation is a process in which electrons are lost whereas reduction is a process in which electrons are accepted/gained.

Rest other statements are correct.

130 (*a*) Statements (a) is incorrect.
It's correct form is as follows :

The given redox reaction for the galvanic cell is,

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

At anode Zn is oxidised to Zn^{2+} ions and at cathode Ag^+ ions are reduced to Ag metal. Thus, galvanic cell for the above redox reaction may be depicted as,

$$Zn \mid Zn^{2+}(aq) \parallel Ag^+(aq) \mid Ag$$

Zn electrode is negatively charged because of the oxidation of Zn to Zn^{2+} ions, electrons are accumulated on zinc electrode.

Rest other statements are correct.

131 (*b*) The reduction potential of mercury is higher than copper and silver. Thus, it can displace both silver and copper from their salt solution.

$$Hg + CuSO_4 \longrightarrow HgSO_4 + Cu$$
$$Hg + AgNO_3 \longrightarrow HgNO_3 + Ag$$

Thus, statement (b) is correct.

132 (*b*) Reduction is the process which involves removal of oxygen or electronegative element from a substance or addition of hydrogen or electropositive element to a substance.

Therefore, statement I and IV are correct and, hence option (b) is correct.

134 (*d*) Statements II and III are correct, while the statement I is incorrect. It's correct form is as follows :

When chlorine gas passes through a concentrated solution of alkali, the following chemical reaction occurs.

$$\overset{-}{Cl_2} + OH^-(aq) \longrightarrow \overset{-1}{Cl^-}(aq) + ClO_3^-$$

Thus, Cl_2 is reduced.

After balancing this equation, we have

$$\underset{3\text{ mol}}{3Cl_2} + \underset{6\text{ mol}}{6OH^-}(aq) \longrightarrow 5Cl^-(aq) + ClO_3^-(aq) + 3H_2O$$

135 (*c*) Statement I, II and IV are correct, while the statement III is incorrect.

It's correct form is as follows :

$$\overset{+2}{Fe}\overset{-2}{SO_4} + K\overset{+5}{Cl}O_3 \longrightarrow K\overset{-1}{Cl} + \overset{+3}{Fe_2}(SO_4)_3^{-2}$$

The change in oxidation state of Fe = 3 − 2 = 1

136 (*a*) Statement I is correct, while the statements II and III are incorrect. Corrected form are as follows :

A more reactive metal can displace the less reactive metal from its salt solution, so in such condition equilibrium lies towards formation product side. Thus, equilibrium greatly favours the formation product of Zn^{2+} and Cu, or Cu^{2+} and Ag. In case of Co and Ni^{2+} neither the reactants, nor the products are greatly favoured.

137 (*a*) When the value of standard reduction potential is negative, the electrode undergoes oxidation and acts as anode. Thus, $Zn^{2+}/Zn; E° = -0.76$ will acts as anode, when connected to standard hydrogen electrode.
Thus, statement I is correct.

138 (*a*) Fluorine is the most electronegative element and hence, attacks water to produce oxygen. That's why the displacement reaction of chlorine, bromine and iodine using fluorine are not generally carried out in aqueous solution.

Thus, both A and R are correct and R is the correct explanation of A.

139 (*a*)

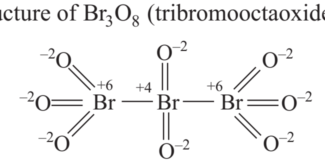
$$2H_2\overset{-1}{O_2} \longrightarrow 2H_2\overset{-2}{O} + \overset{0}{O_2}$$

(Oxidation, Reduction)

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

Thus, both A and R are correct and R is the correct explanation of A.

140 (*b*) The structure of Br_3O_8 (tribromooctaoxide) is



Thus, oxidation state of two corner Br atoms is +6 and of middle one is +4. The difference in oxidation states is due to difference in bonding situations.

$$\text{Average oxidation state} = \frac{+6+4+6}{3} = \frac{16}{3}$$

Thus, both A and R are correct but R is not the correct explanation of A.

141 (*c*) The reaction of potassium permanganate and potassium iodide is as follows :

$$10KI + 2K\overset{+7}{Mn}O_4 + 8H_2SO_4 \longrightarrow 2\overset{+2}{Mn}SO_4 + 6K_2SO_4 + 8H_2O + 5I_2$$

Oxidation state of Mn decreases from +7 to +2.

Thus, A is correct but R is incorrect.

142 (*b*) The electrons are transferred from Zn to Cu^{2+} through the metallic wire which connects the two rods.

While electricity flows through the salt-bridge by migration of ions from one beaker to other.

Thus, both A and R are correct but R is not the correct explanation of A.

143 (*d*) As we know H^+/H_2 couple has zero standard reduction potential so, ions having positive $E°$ value are weaker reducing agent, while ions having negative $E°$ value are stronger reducing agent.

Thus, A is incorrect and R is correct.

144 (*b*) Redox couple is the combination of oxidised and reduced form of a substance. In the representation $E°_{Fe^{3+}/Fe^{2+}}$ and $E°_{Cu^{2+}/Cu}$, Fe^{3+}/Fe^{2+} and Cu^{2+}/Cu are redox couples.

Thus, both A and R are correct but R is not the correct explanation of A.

145 (*b*) The correct match is A $\rightarrow$ 3, B $\rightarrow$ 4, C $\rightarrow$ 1, D $\rightarrow$ 2

Oxidation (Addition of oxygen)

(A) $2Mg + O_2 \longrightarrow 2MgO$

Oxidation (Addition of electronegative element)

(B) $Mg + Cl_2 \longrightarrow MgCl_2$

(C) $2H_2S + O_2 \longrightarrow 2S + 2H_2O$

Reduction (Removal of hydrogen)

(D) $2KI + H_2O + O_3 \longrightarrow 2KOH + I_2 + O_2$

Oxidation (Removal of electropositive element)

147 (*d*) The correct match is A $\rightarrow$ 3, B $\rightarrow$ 2, C $\rightarrow$ 4, D $\rightarrow$ 1

The oxidation number of N-atom in given compounds are shown below :

$$\overset{+3\ \ -2}{N_2O_3},\quad \overset{+1\ +5\ -2}{HNO_3}\quad \overset{+2\ -2}{NO}\quad \overset{-3\ +1\ -2\ +1}{NH_4OH}$$

$$\downarrow +3 \qquad \downarrow +5 \qquad \downarrow +1 \qquad \downarrow -3$$

148 (*c*) The correct match is A $\rightarrow$ 2, B $\rightarrow$ 4, C $\rightarrow$ 1, D $\rightarrow$ 3

The oxidation number of N-atom in given compounds are shown below :

$$\overset{-1+1-2+1}{NH_2OH},\quad \overset{+2\ \ -3}{Mg_3N_2}\quad \overset{+1\ -2}{N_2O}\quad \overset{+5\ -2}{N_2O_5}$$

$$\downarrow -1 \qquad \downarrow -3 \qquad \downarrow -1 \qquad \downarrow +5$$

149 (*a*) The correct match is A $\rightarrow$ 3, B $\rightarrow$ 3, C $\rightarrow$ 1, D $\rightarrow$ 2

The oxidation states of underlined elements in the given compounds are as follows :

$K_4[\underline{Fe}(CN)_6]$, $[\underline{Ni}(CN)_4]^{2-}$, $\underline{N}H_4\underline{N}O_3$, $Ba_2\underline{Xe}O_6$

$4+x+(-1\times6)=0$, $x=+6-4=+2$; $x+(-1\times4)=-2$, $x=-2+4=+2$; -3, $+5$; $+8$

150 (*c*) The correct match is A $\rightarrow$ 1, B $\rightarrow$ 3, C $\rightarrow$ 2, D $\rightarrow$ 4

A. $CH_4(g) + 2O_2 \longrightarrow CO_2(g) + 2H_2O(l)$

It is a combustion reaction.

B. $2NaH(s) \xrightarrow{\Delta} 2Na(s) + H_2(g)$

It is a decomposition reaction.

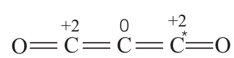
C. $S_8(s) + 12OH^-(aq) \longrightarrow 4S^{2-}(aq) + 2S_2O_3^{2-}(aq) + 6H_2O(l)$

It is a disproportionation reaction as sulphur undergoes oxidation as well as reduction.

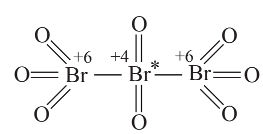
D. $Cr_2O_3(s) + 2Al(s) \longrightarrow Al_2O_3(s) + 2Cr(s)$

It is a displacement reaction, where Al displaces Cr.

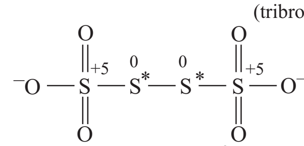
151 (*b*) The correct match is A $\rightarrow$ 3, B $\rightarrow$ 2, C $\rightarrow$ 1



Structure of C_3O_2
(carbon suboxide)



Structure of Br_3O_8
(tribromoctaoxide)



Structure of $S_4O_6^{2-}$
(tetrathionate ion)

152 (*a*) In AgF_2, Ag is in +2 oxidation state. For Ag, this oxidation state is highly unstable so it readily accepts an electron to attain +1 oxidation state which is more stable.

$$Ag^{2+} + e^- \longrightarrow Ag^+$$

That's why, AgF_2 acts as a strong oxidising agent.

153 (*b*) In a reaction, the substance in which oxidation state of an element increases, is oxidised, i.e. acts as reducing agent while that in which oxidation state of an element decreases, is reduced, i.e. acts as oxidising agent.

$$2\overset{+1\ -1}{AgBr} + \overset{-1/3\ +1\ -2}{C_6H_6O_2} \longrightarrow 2\overset{0}{Ag} + 2\overset{+1\ -1}{HBr} + \overset{0\ +1\ -2}{C_6H_4O_2}$$

Substance oxidised (reducing agent)	Substance reduced (oxidising agent)
$C_6H_6O_2(aq)$	$AgBr(s)$

154 (*b*) In (I) and (II) reactions, $AgNO_3$ and $CuSO_4$ act as oxidising agents respectively. They oxidise H_3PO_2 (hypophosphorous acid) to H_3PO_4 (orthophosphoric acid). In reaction (III), $[Ag(NH_3)_2]^+(aq)$ oxidises benzaldehyde to benzoic acid but in reaction (IV), Cu^{2+} do not oxidise benzaldehyde (C_6H_5CHO) to benzoic acid. This indicates that Ag^+ is a stronger oxidising agent than Cu^{2+}.

155 (*d*) The given reaction is,

$$\overset{+3}{(CN)_2}(g) + 2OH^-(aq) \longrightarrow \overset{+2}{CN^-}(aq) + \overset{+4}{CNO^-}(aq) + H_2O(l)$$

Oxidation number decreases

Oxidation number increases

The following information we can draw from the above reaction.

(a) Decomposition of cyanogen into cyanide ion (CN^-) and cyanate ion (CNO^-) occurs in basic medium.

(b) Cyanogen $(CN)_2$ acts as both reducing agent as well as oxidising agent.

(c) The reaction is an example of disproportionation reaction.

156 (*d*) The first step of Ostwald's process is as follows :

$$\underset{\substack{4\times17\\=68\text{ g}}}{4NH_3(g)} + \underset{\substack{5\times32\\=160\text{ g}}}{5O_2(g)} \xrightarrow[\text{Pt}]{1100\text{ K}} \underset{\substack{4\times30\\=120\text{ g}}}{4NO(g)} + 6H_2O(g)$$

68 g NH_3 reacts with 160 g O_2

$$1\text{ g } NH_3 \text{ reacts with } \frac{160\times1}{68}\text{ g } O_2$$

$$\therefore\ 10\text{ g } NH_3 \text{ will react with } \frac{160\times10}{68} = 23.5\text{ g } O_2$$

But available amount of O_2 is 20.0 g which is less than the amount which is required to react with 10 g NH_3.

So,O_2 is the limiting reagent and it limits the amount of NO produced.

From the above balanced equation, 160 g of O_2 produces 120 g NO.

$$1\text{ g of } O_2 \text{ produces } \frac{120\times1}{160}\text{ g NO}$$

$$\therefore\ 20\text{ g of } O_2 \text{ will produce } \frac{120\times1\times20}{160} = 15\text{ g NO}$$

157 (*a*) An aqueous solution of $AgNO_3$ with silver electrode, when undergoes electrolysis, two oxidation and two reduction half-reactions must be considered.

Oxidation (at anode)

(A) $Ag(s) \longrightarrow Ag^+(aq) + e^-;\ E° = -0.80$ V

(B) $2H_2O(l) \longrightarrow O_2(g) + 4H^+(aq) + 4e^-;$
$E° = -1.23$ V

Reduction (at cathode)

(C) $Ag^+(aq) + e^- \longrightarrow Ag(s);\quad E° = +0.80$ V

(D) $2H_2O(l) + 2e^- \longrightarrow H_2(g) + 2OH^-(aq);$
$E° = -0.83$V

By $E°$ values of (A) and (B), it appears that at anode silver of silver anode gets oxidised more readily because oxidation potential of Ag is greater than that of H_2O molecule. Similarly by $E°$ values of (C) and (D), it appears that at cathode reduction potential of Ag^+ ions is higher than that of H_2O molecules.

Therefore, on electrolysis of *aq*. $AgNO_3$ solution with silver electrodes, Ag from silver anode dissolves while $Ag^+(aq)$ ions present in the solution get reduced and deposited at cathode.

158 (*c*) $E°_{Al^{3+}/Al} = -1.66$ V; $\quad E°_{Cu^{2+}/Cu} = +0.34$ V

$E°_{Fe^{2+}/Fe} = -0.44$ V; $\quad E°_{Mg^{2+}/Mg} = -2.36$ V

and $E°_{Zn^{2+}/Zn} = -0.76$ V

A metal with more negative value of $E°_{red}$ is a stronger reducing agent than those which have less negative or positive value of $E°_{red}$. Therefore, Mg can displace all the given metals from their aqueous salt solutions.

Al can displace all metals except Mg from their aqueous salt solutions. Zinc can displace Fe and Cu from their aqueous salt solutions and Fe can only displace Cu from its aqueous salt solution. Hence, the order in which they can displace each other from the solution of their salts is as follows :

Mg, Al, Zn, Fe, Cu

159 (*c*) More negative $E°$ value means that the redox couple is a stronger reducing agent. The arrangement of metals in their increasing order of reducing power is as follows:

$$Ag < Hg < Cr < Mg < K.$$

160 (*d*) Reaction $\overset{+2}{Ba}\overset{-1}{Cl_2} + \overset{+1}{H_2}\overset{+6}{S}\overset{-2}{O_4} \longrightarrow \overset{+2}{Ba}\overset{+6}{S}\overset{-2}{O_4} + 2\overset{+1}{H}\overset{-1}{Cl}$

Option (d) is not an example of redox reaction.

It is because in this reaction there is no change in oxidation number of the interacting species.

161 (*d*) Since, $E°$ value of the redox couple Ag^+/Ag is the most positive, i.e. + 0.80 V, therefore, Ag^+ is the strongest oxidising agent amongst given options.

162 (*d*) The $E°$ values show that copper will reduce Br_2, if the $E°$ of the following redox reaction is positive.

$$Cu + Br_2 \longrightarrow CuBr_2$$

Now, $\quad Cu \longrightarrow Cu^{2+} + 2e^-;\ E° = -0.34$V

$$Br_2 + 2e^- \longrightarrow 2Br^-;\ E = +1.09\text{V}$$

$$Cu + Br_2 \longrightarrow CuBr_2;\ E° = +0.75\text{ V}$$

Since, $E°$ of this reaction is positive, therefore, Cu can reduce Br_2. While other reaction will give negative value of E°_{cell}.

163 (*a*) $2\overset{+2}{S_2}\overset{-2}{O_3^{2-}}(aq) + I_2^0(s) \longrightarrow \overset{2.5}{S_4}O_6^{2-}(aq) + 2I^-(aq)$

$\overset{+2}{S_2}\overset{-2}{O_3^{2-}}(aq) + 2\overset{0}{Br_2}(l) + 5H_2O(l) \longrightarrow$

$$2\overset{+6}{S}\overset{2-}{O_4^{2-}}(aq) + 4Br^-(aq) + 10H^+(aq)$$

Bromine being stronger oxidising agent than iodine, oxidises S (in +2 oxidation state) of $S_2O_3^{2-}$ ion to

S (in +6 oxidation state) of SO_4^{2-} ion, whereas

I_2 oxidises S from +2 oxidation state of $S_2O_3^{2-}$ to only +2.5 oxidation state of $S_4O_6^{2-}$ ion.

164 (*a*) Oxidation number of hydrogen is always +1 is a wrong since, it is +1 in hydrogen halides, −1 in hydrides and zero in H_2 molecule.

165 (*b*) NH_4NO_3 is actually NH_4^+ and NO_3^-. It is an ionic compound. The oxidation number of nitrogen in the two species is different.

Let, oxidation number of N in NH_4^+ is *x*.

$\Rightarrow x + (4 \times 1) = +1$ or $x + 4 = +1$ or $x = -3$

Let oxidation number of N in NO_3^- is *x*.

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

166 (*a*) The oxidation number of Cr, Cl and Mn of each species in the four set of ions, is

$$\overset{+3}{Cr}O_2^-, \overset{+5}{Cl}O_3^-, \overset{+6}{Cr}O_4^{2-}, \overset{+7}{Mn}O_4^-$$

Thus, the correct arrangements representing increasing oxidation number of the central atom is

$$CrO_2^-, ClO_3^-, CrO_4^{2-}, MnO_4^-$$

167 (*d*) Highest oxidation number of any transition element = $(n-1)$ *d*-electrons + *ns*-electrons. Therefore, larger the number of unpaired electrons in the 3*d*-orbitals, higher is the oxidation number.

(a) $3d^1 4s^2 = 3$ (b) $3d^3 4s^2 = 3 + 2 = 5$

(c) $3d^5 4s^1 = 5 + 1 = 6$ and (d) $3d^5 4s^2 = 5 + 2 = 7$

168 (*d*) (a) $\overset{-4}{C}\overset{+1}{H_4} + 2\overset{0}{O_2} \longrightarrow \overset{+4}{C}\overset{-2}{O_2} + 2\overset{+1}{H_2}\overset{-2}{O}$

(b) $\overset{-4}{C}\overset{+1}{H_4} + 4\overset{0}{Cl_2} \longrightarrow \overset{+4}{C}\overset{-1}{Cl_4} + 4\overset{+1}{H}\overset{-1}{Cl}$

(c) $2\overset{0}{F_2} + 2\overset{-2}{O}\overset{+1}{H} \longrightarrow 2\overset{-1}{F^-} + \overset{+2}{O}\overset{-1}{F_2} + \overset{+1}{H_2}\overset{-2}{O}$

(d) $2\overset{+4}{N}\overset{-2}{O_2} + 2\overset{-2}{O}\overset{+1}{H^-} \longrightarrow \overset{+3}{N}\overset{-2}{O_2} + \overset{+5}{N}\overset{-2}{O_3^-} + \overset{+1}{H_2}\overset{-2}{O}$

Thus, in reaction (d), N is both oxidised as well as reduced since the ON (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^-.