Blue Print (As per PU Board)

Topic	1 mark questions	2 marks questions	3 marks questions	5 marks questions	Total Marks
Redox Reactions	1	–	1	–	4

One mark questions

1. **Define 'oxidation' in terms of electron transfer.**
   Answer: Loss of electron(s) by any species is called oxidation.
2. **Which is the most powerful oxidizing agent?**
   Answer: Fluorine (F_2).
3. **What is the oxidation state of hydrogen in hydrides?**
   Answer: In Hydrides, hydrogen has an oxidation state of -1
4. **What is electrode potential?**
   Answer: The potential attained by a metal in contact with a solution containing its own ions is called electrode potential.
5. **What is the oxidation state of oxygen in OF_2?.**
   Answer: +2

Two marks questions

6. **Justify the reaction:** $\mathrm{H_2S + Cl_2 \rightarrow 2HCl + S}$ **is a redox reaction.**
   Answer: $\mathrm{H_2S^{(-2)} + Cl_2^{(0)} \rightarrow 2HCl^{(-1)} + S(0)}$
   The O.N. of S increases from -2 to 0. So it is undergoing oxidation.
   The O.N. of $\mathrm{Cl_2}$ decreases from 0 to -1. So it is undergoing reduction.
   Therefore it is a redox reaction.
7. **Define oxidation and reduction in terms of oxidation number.**
   Answer: In terms of oxidation number,
   Oxidation: An increase in the oxidation number of an element in a given substance.
   Reduction: A decrease in the oxidation number of an element in a given substance.
8. **What is an electrochemical series?**
   Answer: A series of electrode potential values arranged in the increasing or decreasing order constitute an electrochemical series.
9. **Give an example of a redox disproportionation reaction. Mention the species that undergo oxidation and reduction.**
   Answer: Example for Redox disproportionation reaction:
   $+1-1 \qquad +1-2 \qquad 0$
   $2H_2O_2(aq) \rightarrow H_2O(I) + O_2(g)$
   In this reaction, the O.N. of 'O' increases from -1 to 0 as well as decreases from -1 to -2.
   So oxygen is undergoing both oxidation and reduction (disproportionation).

Five marks questions

10. **When blue coloured solution of copper sulphate is stirred with a zinc rod, the blue colour of the solution fades off and the zinc rod is coated with reddish copper metal. Write the chemical reaction taking place in the above observation and identify the species undergoing oxidation and reduction.**
    Answer:

reduction

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

Oxidation

In this reaction, Zn loses $2e^-$ to Cu and hence is undergoing oxidation; Cu^{2+} is undergoing reduction to Cu.

11. **Balance the following equations by the oxidation number method.**

(i) $Fe^{2+} + H^+ + Cr_2O_7^{2-} \rightarrow Cr^{3+} + Fe^{3+} + H_2O$

(ii) $MnO_4^-(aq) + I^-(aq) \rightarrow MnO_2(s) + I_2(s)$ **(in basic medium)**

(iii) $MnO_4^-(aq) + SO_2(g) \rightarrow Mn^{2+}(aq) + HSO_4^-(aq)$ **(in acidic solution)**

(iv) $H_2O_2(aq) + Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + H_2O(I)$ **(in acidic solution)**

(v) $Cr_2O_7^{2-} + SO_2(g) \rightarrow Cr^{3+}(ap) + SO_4^{2-}(aq)$ **(in acidic solution)**

Answer: (i) $Fe^{2+} + H^+ + Cr_2O_7^{2-} \rightarrow Cr^{3+} + Fe^{3+} + H_2O$

Step 1: Write skeletal equation with O.N of each element.

Oxidation

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

reduction

Step 2: Multiply Cr^{3+} by 2 and Fe^{2+} by 6 to equalize the oxidation numbers on either side of the equation.

$$\overset{2+}{6Fe^{2+}} + \overset{1+}{H^+} + \overset{6+}{Cr_2O_7^{2-}} \rightarrow \overset{2\times3+}{2\,Cr^{3+}} + \overset{3+}{6\,Fe^{3+}} + H_2O$$

Step 3: Now, balance O atoms on RHS by adding $7H_2O$

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

Step 4: Finally balance H atoms by adding $14H^+$ on LHS to get a balanced equation as:

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

(ii) $MnO_4^-(aq) + I^-(aq) \rightarrow MnO_2(s) + I_2(s)$ (in basic medium)

Step 1: Write skeletal equation with O.N of each element Undergoing change in oxidation number

Oxidation

$$\overset{7+}{MnO_4^-}(aq) + \overset{1-}{I^-}(aq) \rightarrow \overset{6+}{MnO_2}(s) + \overset{0}{I_2}(s)$$

reduction

Step 2: Multiply I-by 6 and MnO_4^- by 2 to equalize the oxidation numbers on either side of the equation.

$$\overset{2\times7+}{2MnO_4^-}(aq) + \overset{6\times(1-)}{6I^-}(aq) \rightarrow \overset{2\times(4+)}{2MnO_2}(s) + \overset{0}{3I_2}(s)$$

Step 3 : Now, and 8 OH^- on RHS to balance –ve charges on either side.

$$2\,MnO_4^-(aq) + 6I^-(aq) \rightarrow 2MnO_2(s) + 3I_2(s) + 8OH^-$$

Step 4: Finally balance H and O atoms by adding $4H_2O$ on LHS to get a balanced equation as:

$$2MnO_4^-(aq) + 6I^-(aq) + 4H_2O \rightarrow 2MnO_2(s) + 3\,I_2(s) + 80H^-$$

(iii) $MnO_4^-(aq) + SO_2(g) \rightarrow Mn^{2+}(aq) + HSO_4^-(aq)$ (in acidic solution)

Step 1: Write skeletal equation with O.N of each element undergoing change in oxidation number.

Oxidation - $2e^-$

$$\overset{7+}{MnO_4^-}(aq) + \overset{4+}{SO_2}(g) \rightarrow \overset{2+}{Mn^{2+}}(aq) + \overset{6+}{HSO_4^-}(aq)$$

reduction - $5e^-$

Step 2: Multiple SO_2 by 5 and MnO_4^- by 2 to balance +ve charges on both sides.

$$\begin{array}{cccc} 2\times(7+) & 5\times(4+) & 2\times(2+) & 5\times(6+) \end{array}$$

$$2MnO_4^-(aq) + 5SO_2(g) \rightarrow 2Mn^{2+}(aq) + 5HSO_4^-(aq)$$

Step 3: Now, and 2H2O and $H+$ on LHS to balance oxygen atoms

$$2MnO_4^-(aq) + 5\,SO_2(g) + 2H2O \rightarrow 2Mn^{2+}(aq) + 5\,HSO_4^-(aq)$$

Step 4: Finally add H^+ on LHS to get a balanced equation as

$$2\,MnO_4^-(aq) + 5SO_2(g) + 2H_2O + H^+ \rightarrow 2Mn^{2+}(aq) + 5HSO_4^-(aq)$$

(iv) $H_2O_2(aq) + Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + H_2O(I)$ (in acidic solution)

Step 1 : Write skeletal equation with O.N of each element undergoing change in oxidation number

Oxidation - $1e^-$

$$\overset{1-}{H_2O_2}(aq) + \overset{2+}{Fe^{2+}}(aq) \rightarrow \overset{3+}{Fe^{3+}}(aq) + \overset{2-}{H_2O}(l))$$

reduction-2 x 1e-

Step 2 : Since the number of charges on both sides are not equal, $2Fe^{2+}$ on LHS and $2Fe^{3+}$ on RHS

$$\begin{array}{cccc} 2\times(1-) & 2\times(2+) & 2\times(3+) & (2-) \end{array}$$

$$H_2O_2(aq) + 2Fe^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + H_2O(1))$$

Step 3: Now, put $2H_2O$ to balance 'O' atoms.

$$H_2O_2(aq) + Fe^{2+}(aq) + 2H^+(aq) \rightarrow 2Fe^{3+}(aq) + 2H_2O(1))$$

Step 4: Finally and $2\,H^+$ on LHS to get a balanced equation as:

$$H_2O_2(aq) + 2\,Fe^{2+}(aq) + 2H^+(aq) \rightarrow 2Fe^{3+}(aq) + H_2O(1))$$

(v) $Cr_2O_7^{2-} + SO_2(g) \rightarrow Cr^{3+}(ap) + SO_4^{2-}(aq)$ (in acidic solution)

Step 1: Write skeletal equation with O.N of each element Undergoing change in oxidation number.

Oxidation-2e-

$$\overset{2\times(6+)}{Cr_2O_7^{2-}} + \overset{4+}{SO_2}(g) \rightarrow \overset{3+}{Cr^{3+}}(aq) + \overset{6+}{SO_4^{2-}}(aq)$$

reduction-3e-

Step 2 : Multiply SO_2 by 3 and $Cr3^{+}$ by 2 on RHS

$Cr_2O_7^{2-} + 3\,SO_2(g) \rightarrow 2Cr^{3+}(aq) + 3\,SO_4^{2-}(aq)$

Step 3: Balance charges by adding $2H^{+}$ on LHS

$Cr_2O_7^{2-} + 3\,SO_2(g) + 2H^{+} \rightarrow 2Cr^{3+}(aq) + 3SO_4^{2-}(aq)$

Step 4: Finally add H_2O on RHS to get a balanced equation as:

$Cr_2O_7^{2+} + SO_2(g) \rightarrow 2Cr^{3+}(aq) + SO_4^{2-}(aq) + H_2O(I)$

12. **Balance the following equations by half reaction method (ion-electron method).**

(i) $MnO_4^{-}(aq) + SO_2(g) \rightarrow Mn^{2+}(aq) + HSO_4^{-}(aq)$ **(in acidic solution)**

(ii) $MnO_4^{-}(aq) + I_2(aq) \rightarrow MnO_2 + I_2(s)$ **(in basic solution)**

(iii) $H_2O_2(aq) + Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + H_2O(I)$ **(in acidic solution)**

(iv) $Cr_2O_7^{2-} + SO_2(g) \rightarrow Cr^{3+}(aq) + SO_4^{2-}$ **(aq) (in acidic solution)**

(v) $MnO_4^{-} + C_2O_4^{2-} \rightarrow Mn^{2+} + CO_2$ **in acidic solution**

Answer:

(i) $MnO_4^{-}(aq) + SO_2(g) \rightarrow Mn^{2+}(aq) + HSO_4^{-}(aq)$ (in acidic solution)

Step 1: Assign O.N. to the atom undergoing oxidation / reduction.

oxidation-2e-

$MnO_4^{-}(aq) + SO_2(g) \rightarrow Mn^{2+}(aq) + HSO_4^{-}(aq)$

reduction-5e-

Step 2: Write out oxidation and reduction separately and balance the atoms other than H and O.

Oxidation half reaction $SO_2 \rightarrow HSO_4^{-}$

Reduction half reaction: $MnO_4^{-} \rightarrow Mn^{2+}$

Step 3: Multiply the oxidation reaction with the extent of reduction and reduction reaction by the extent of oxidation and add.

Oxidation half reaction: $\left[SO_2 \rightarrow HSO_4^{-}\right] \times 5$

Reduction half reaction: $\left[MnO_4^{-} \rightarrow Mn^{2+}\right] \times 2$

$$2MnO_4^{-} + 5SO_2 \rightarrow 2Mn^{2+} + 5HSO_4^{-}$$

Step 4: Add H^{+} and $2H_2O$ on LHS to balance H and O atoms in the acid medium to get a balanced equation.

$2MnO_4^{-} + 5\,SO_2 + H^{+} + 2\,H_2O \rightarrow 2Mn^{2+} + 5HSO_4^{-}$

(ii) $MnO_4^{-}(aq) + I_2(aq) \rightarrow MnO_2 + I_2(s)$ (in basic solution)

Step 1: Assign O.N. to the atoms undergoing oxidation / reduction.

Oxidation - $2e^{-}$

$MnO_4^{-}(aq) + I^{-}(aq) \rightarrow MnO_2(s) + I_2(s)$

Step 2: Write out oxidation and reduction separately and balance the atoms other than H and O.

Oxidation half reaction $2I^{-} \rightarrow I_2$

Reduction half reaction: $MnO_4^{-} \rightarrow MnO_2$

Step 3: Multiply the oxidation reaction with the extent of reduction and reduction reaction by the extent of oxidation and add.

Oxidation half reaction: $\left[2\ I^{-} \rightarrow I_2\right] \times 3$

Reduction half reaction: $\left[MnO_4^{-} \rightarrow MnO_2\right] \times 2$

$$2MnO_4^{-} + 6I^{-} \rightarrow MnO_2 + 3I_2$$

Step 4: Add $4\ OH^{-}$ on RHS and $2H_2O$ on LHS to balance H and O atoms in the basic medium to get a balanced equation.

$2\ MnO_4^{-} + 6I^{-} + 4H_2O \rightarrow 2\ MnO_2 + 2\ I_2 + 8\ OH^{-}$

(iii) $H_2O_2(aq) + Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + H_2O(I)$ (in acidic solution)

Step 1 : Assign O.N. to the atoms undergoing oxidation / reduction.

Oxidation - $1e^{-}$

$$H_2O_2(aq) + Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + H_2O(I)$$

reduction - $2e^{-}$

Step 2: Write out oxidation and reduction separately and balance the atoms other than H and O.

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

Reduction half reaction: $H_2O_2 \rightarrow H_2O$

Step 3: Multiply the oxidation reaction with the extent of reduction and reduction reaction by the extent of oxidation and add.

Oxidation half reaction: $\left[Fe^{2+} \rightarrow Fe^{3+}\right] \times 2$

Reduction half reaction: $\left[H_2O_2 \rightarrow H_2O\right] \times 1$

$$2Fe^{2+} + H_2O_2 \rightarrow 2Fe^{3+} + H_2O$$

Step 4: Add $2H^{+}$ on LHS and H_2O on RHS to balance H and O atoms in the acid medium to get a balanced equation.

$2Fe^{2+} + H_2O_2 + 2H^{+} \rightarrow 2Fe^{3+} + 2\ H_2O$

(iv) $Cr_2O_7^{2-} + SO_2(g) \rightarrow Cr^{3+}(aq) + SO_4^{2-}$ (aq) (in acidic solution)

Step 1 : Assign O.N. to the atoms undergoing oxidation / reduction.

Oxidation - $2e^{-}$

$$Cr_2O_1^{2-} + SO_2(g) \rightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$$

reduction - $2 \times 3\ e^{-}$

Step 2: Write out oxidation and reduction separately and balance the atoms other than H and O.

Oxidation half reaction: $SO_2 \rightarrow SO_4^{2-}$

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

Step 3: Multiply the oxidation reaction with the extent of reduction and reduction reaction by the extent of oxidation and add.

Oxidation half reaction: $\left[SO_2 \rightarrow SO_4^{2-}\right] \times 6$

Reduction half reaction: $\left[Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+}\right] \times 2$

$$2Cr_2O_7^{2-} + 6\,SO_2 \rightarrow 4Cr^{3+} + 6SO_4^{2-}$$

Step 4: Add H^+ and $2H_2O$ on LHS to balance H and O atoms in the acid medium to get a balanced equation.

$2Cr_2O_7^{2-}(aq) + 6\,SO_2(g) + 4H^+ \rightarrow 4\,Cr^{3+}(aq) + 6\,SO_4^{2-}(aq) + 2H_2O$

Or $Cr_2O_7^{2-}(aq) + 3\,SO_2(g) + 2H^+ \rightarrow 2\,Cr^{3+}(aq) + 3SO_4^{2-}(aq) + H_2O$

(v) $MnO_4^- + C_2O_4^{2-} \rightarrow Mn^{2+} + CO_2$ in acidic solution

Step 1: Assign O.N. to the atoms undergoing oxidation / reduction.

Oxidation - $2 \times 1e^-$

$$MnO_4^- + C_2O_4^{2-} \rightarrow Mn^{2+} + CO_2$$

reduction - $5e^-$

Step 2: Write out oxidation and reduction separately and balance the atoms other than H and O.

Oxidation half reaction: $C_2O_4^{2-} \rightarrow 2CO_2$

Reduction half reaction: $MnO_4^- \rightarrow Mn^{2+}$

Step 3: Multiply the oxidation reaction with the extent of reduction and reduction reaction by the extent of oxidation and add.

Oxidation half reaction: $\left[C_2O_4^{2-} \rightarrow 2CO_2\right] \times 5$

Reduction half reaction: $\left[MnO_4^- \rightarrow Mn^{2+}\right] \times 2$

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

Step 4: Add required number H^+ on LHS and H_2O on RHS to balance H and O atoms in the acid medium to get a balanced equation.

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