



## 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**

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

**Two marks questions**

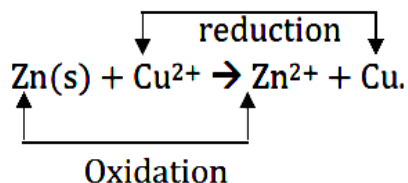
- Justify the reaction:  $H_2S + Cl_2 \rightarrow 2HCl + S$  is a redox reaction.**  
Answer:  $H_2S^{(-2)} + Cl_2^{(0)} \rightarrow 2HCl^{(-1)} + S^{(0)}$   
The O.N. of S increases from -2 to 0. So it is undergoing oxidation.  
The O.N. of  $Cl_2$  decreases from 0 to -1. So it is undergoing reduction.  
Therefore it is a redox reaction.
- 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.
- What is an electrochemical series?**  
Answer: A series of electrode potential values arranged in the increasing or decreasing order constitute an electrochemical series.
- Give an example of a redox disproportionation reaction. Mention the species that undergo oxidation and reduction.**  
Answer: Example for Redox disproportionation reaction:  

$$+1-1 \quad +1-2 \quad 0$$

$$2H_2O_2(aq) \rightarrow H_2O(l) + O_2(g)$$
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**

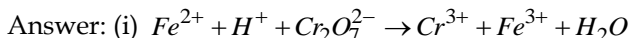
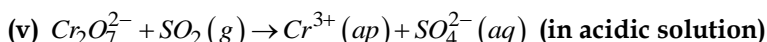
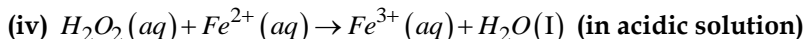
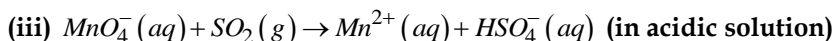
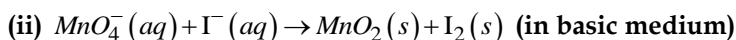
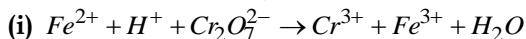
- When blue coloured solution of copper sulphate is stirred with a zinc rod, the blue colour of the solution fades off and the zinc rod is coated with reddish copper metal. Write the chemical reaction taking place in the above observation and identify the species undergoing oxidation and reduction.**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



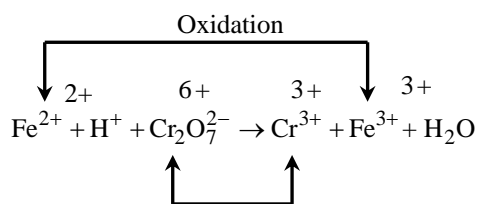


In this reaction, Zn loses  $2e^-$  to Cu and hence is undergoing oxidation;  $Cu^{2+}$  is undergoing reduction to Cu.

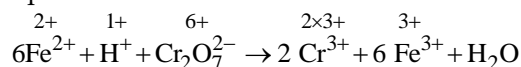
11. Balance the following equations by the oxidation number method.



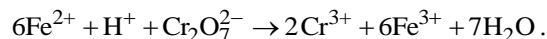
Step 1: Write skeletal equation with O.N of each element.



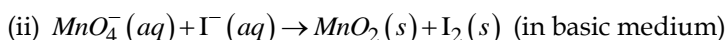
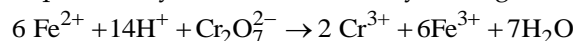
Step 2: Multiply  $Cr^{3+}$  by 2 and  $Fe^{2+}$  by 6 to equalize the oxidation numbers on either side of the equation.



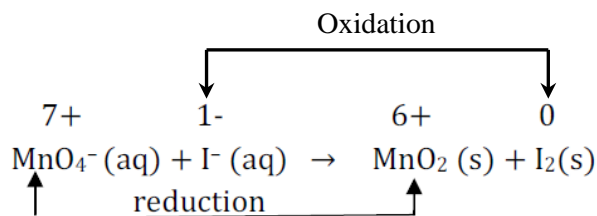
Step 3: Now, balance O atoms on RHS by adding  $7H_2O$



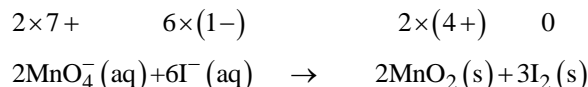
Step 4: Finally balance H atoms by adding  $14H^+$  on LHS to get a balanced equation as:



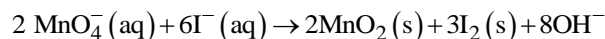
Step 1: Write skeletal equation with O.N of each element Undergoing change in oxidation number



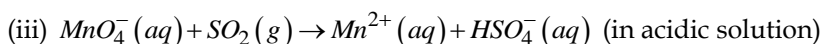
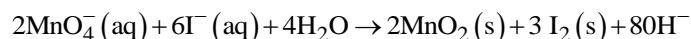
Step 2: Multiply I-by 6 and  $MnO_4^-$  by 2 to equalize the oxidation numbers on either side of the equation.



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

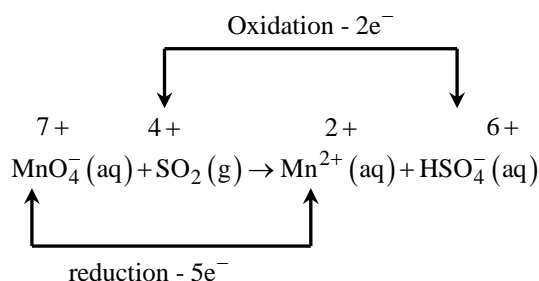


Step 4: Finally balance H and O atoms by adding  $4H_2O$  on LHS to get a balanced equation as:

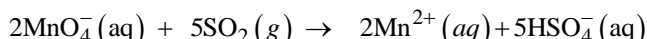




Step 1: Write skeletal equation with O.N of each element undergoing change in oxidation number.



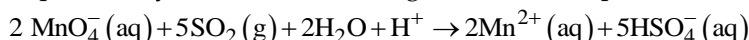
Step 2: Multiple  $\text{SO}_2$  by 5 and  $\text{MnO}_4^-$  by 2 to balance +ve charges on both sides.



Step 3: Now, add  $2\text{H}_2\text{O}$  and  $\text{H}^+$  on LHS to balance oxygen atoms

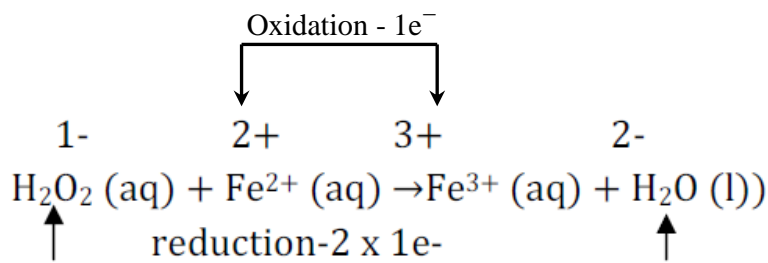


Step 4: Finally add  $\text{H}^+$  on LHS to get a balanced equation as

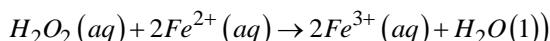
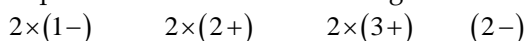


(iv)  $\text{H}_2\text{O}_2 (\text{aq}) + \text{Fe}^{2+} (\text{aq}) \rightarrow \text{Fe}^{3+} (\text{aq}) + \text{H}_2\text{O} (\text{l})$  (in acidic solution)

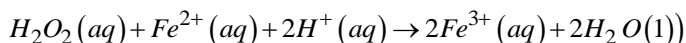
Step 1 : Write skeletal equation with O.N of each element undergoing change in oxidation number



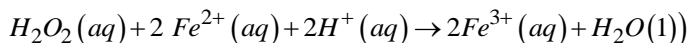
Step 2 : Since the number of charges on both sides are not equal,  $2\text{Fe}^{2+}$  on LHS and  $2\text{Fe}^{3+}$  on RHS



Step 3: Now, put  $2\text{H}_2\text{O}$  to balance 'O' atoms.

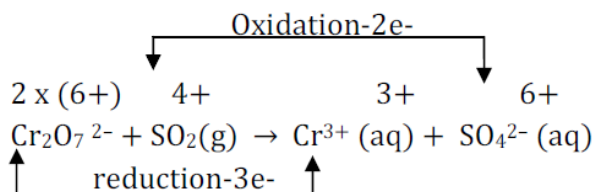


Step 4: Finally add  $2\text{H}^+$  on LHS to get a balanced equation as:



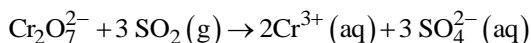
(v)  $\text{Cr}_2\text{O}_7^{2-} + \text{SO}_2 (\text{g}) \rightarrow \text{Cr}^{3+} (\text{aq}) + \text{SO}_4^{2-} (\text{aq})$  (in acidic solution)

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

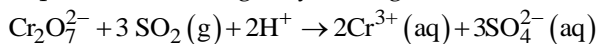




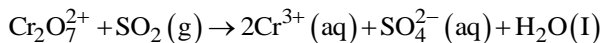
Step 2 : Multiply  $\text{SO}_2$  by 3 and  $\text{Cr}^{3+}$  by 2 on RHS



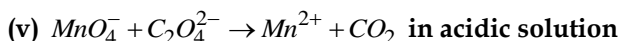
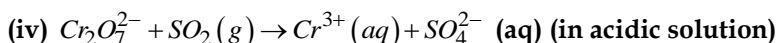
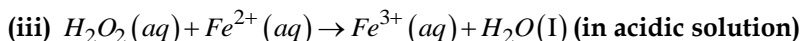
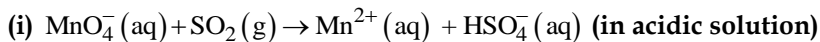
Step 3: Balance charges by adding  $2\text{H}^+$  on LHS



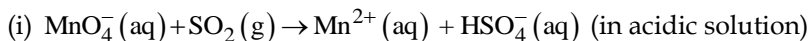
Step 4: Finally add  $\text{H}_2\text{O}$  on RHS to get a balanced equation as:



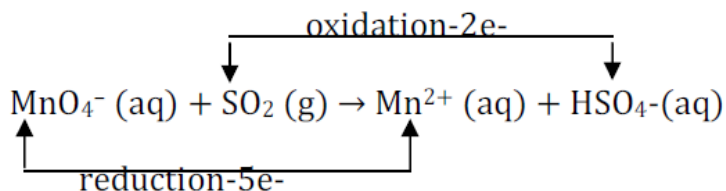
12. Balance the following equations by half reaction method (ion-electron method).



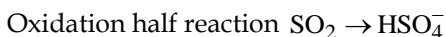
Answer:



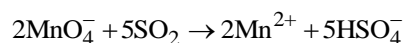
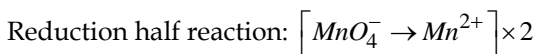
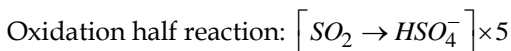
Step 1: Assign O.N. to the atom undergoing oxidation / reduction.



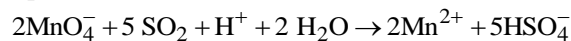
Step 2: Write out oxidation and reduction separately and balance the atoms other than H and O.



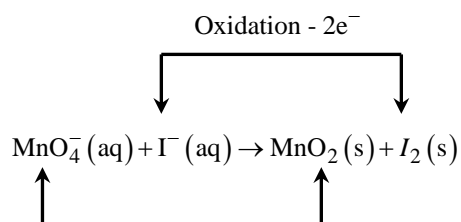
Step 3: Multiply the oxidation reaction with the extent of reduction and reduction reaction by the extent of oxidation and add.



Step 4: Add  $\text{H}^+$  and  $2\text{H}_2\text{O}$  on LHS to balance H and O atoms in the acid medium to get a balanced equation.

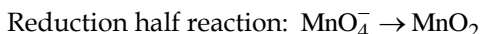
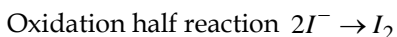


Step 1: Assign O.N. to the atoms undergoing oxidation / reduction.

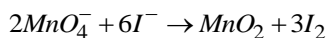
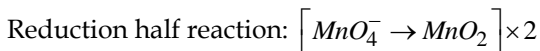
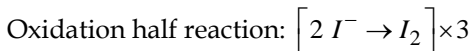




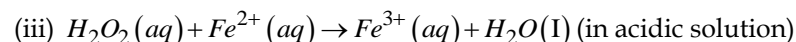
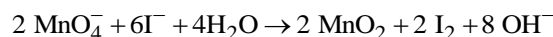
Step 2: Write out oxidation and reduction separately and balance the atoms other than H and O.



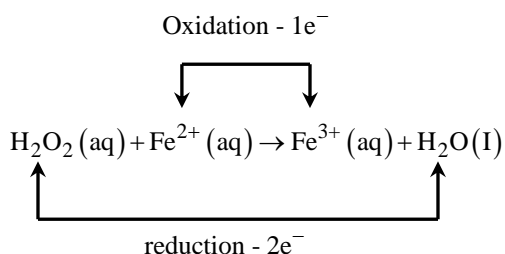
Step 3: Multiply the oxidation reaction with the extent of reduction and reduction reaction by the extent of oxidation and add.



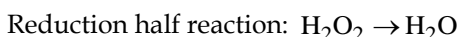
Step 4: Add 4 OH<sup>-</sup> on RHS and 2H<sub>2</sub>O on LHS to balance H and O atoms in the basic medium to get a balanced equation.



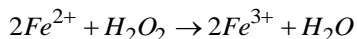
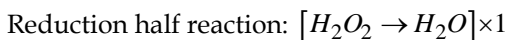
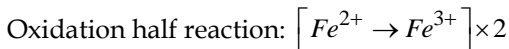
Step 1 : Assign O.N. to the atoms undergoing oxidation / reduction.



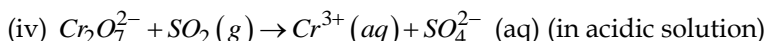
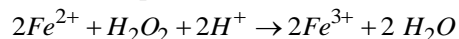
Step 2: Write out oxidation and reduction separately and balance the atoms other than H and O.



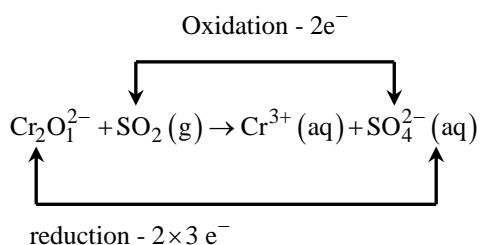
Step 3: Multiply the oxidation reaction with the extent of reduction and reduction reaction by the extent of oxidation and add.



Step 4: Add 2H<sup>+</sup> on LHS and H<sub>2</sub>O on RHS to balance H and O atoms in the acid medium to get a balanced equation.

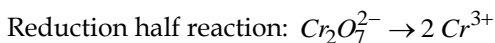


Step 1 : Assign O.N. to the atoms undergoing oxidation / reduction.

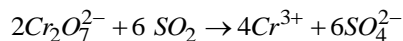
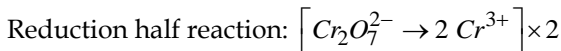
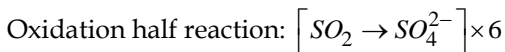


Step 2: Write out oxidation and reduction separately and balance the atoms other than H and O.

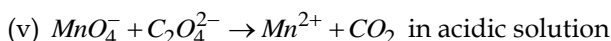
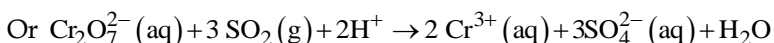
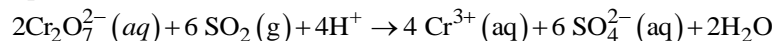




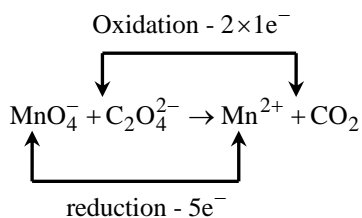
Step 3: Multiply the oxidation reaction with the extent of reduction and reduction reaction by the extent of oxidation and add.



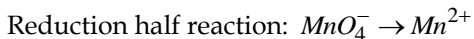
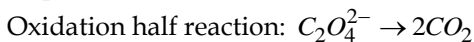
Step 4: Add  $H^+$  and  $2H_2O$  on LHS to balance H and O atoms in the acid medium to get a balanced equation.



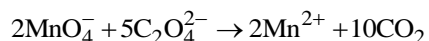
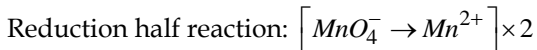
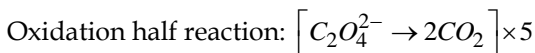
Step 1: Assign O.N. to the atoms undergoing oxidation / reduction.



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



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



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.

