

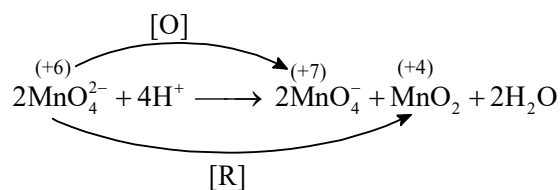
## HINTS & SOLUTIONS

### Redox Reactions

### [Set-2]

#### SECTION-A

1. (iii): In  $\text{HOClO}_2$  O.S of Cl is +5 and in  $\text{ClO}_4^-$ , O.S of Cl is +7
2. (ii):



3. (i):  $\text{PCl}_5 + \text{H}_2\text{O} \longrightarrow \text{POCl}_3 + 2\text{HCl}$

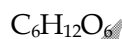
No change in the O.S of phosphorus

4. (i):  $\text{N}_2\text{H}_4 + 2\text{H}_2\text{O}_2 \longrightarrow \text{N}_2 + 4\text{H}_2\text{O}$

O.S of nitrogen is changing from -2 to zero, loss of electrons.

Hence,  $\text{N}_2\text{H}_4$  acting as a reducing agent.

5. (iv):
6. (ii): Oxidation state of carbon in glucose is zero.

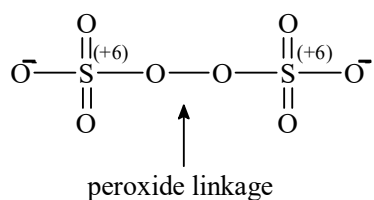


$$6x + 12(+1) + 6(-2) = 0$$

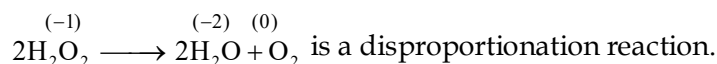
$$x = 0$$

Glucose is a covalent compound.

7. (i):  $\text{S}_2\text{O}_8^{2-}$  ion has a structure



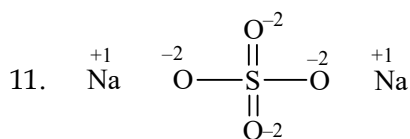
8. (iv): Assertion is wrong



9. (ii):  $\overset{+2}{\text{CO}} + \overset{+4}{\text{H}_2\text{O}} \longrightarrow \overset{+4}{\text{H}_2} + \overset{+4}{\text{CO}_2}$

CO acts as a reducing agent.

10. (iv):  $\overset{3+}{\text{Fe}} \xrightarrow{\quad} \overset{2-}{\text{SO}_4} \longrightarrow \text{Fe}_2(\text{SO}_4)_3$



12. (i)  $\text{MgO} + \text{C} \rightarrow \text{Mg} + \text{CO}$  is a redox reaction.  $\text{MgO} \rightarrow \text{Mg}$  is a loss of oxygen and is called **reduction** and  $\text{C} \rightarrow \text{CO}$  is a gain of oxygen, called **oxidation**.

(ii)  $\text{CO}_2 + \text{C} \rightarrow 2\text{CO}$  is also a redox reaction.

$\text{CO}_2 \rightarrow \text{CO}$  is a loss of oxygen and is called reduction and  $\text{C} \rightarrow \text{CO}$  is a gain of oxygen, called oxidation

**OR**

(i)  $\text{Na}_2\text{S}_4\text{O}_6$

Let the O.S of sulphur be x, then

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

$$x = \frac{10}{4} = 2.5$$

(ii) Let the O.S of chromium be x, then

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

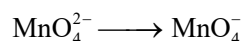
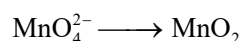
$$x - 8 + 4 = -1$$

$$x = +3$$

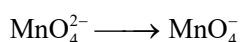
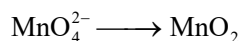
13. The skeletal equation is



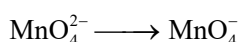
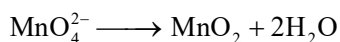
(a) Split the redox reaction into two half reactions,



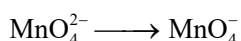
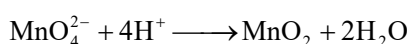
(b) Balance all atoms except H and O



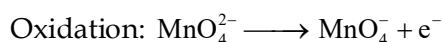
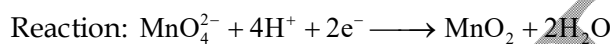
(c) Balance O atoms by adding H<sub>2</sub>O



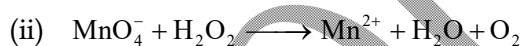
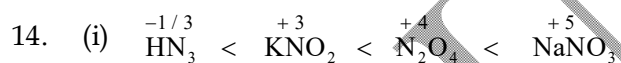
(d) Balance H atoms by adding H<sup>+</sup> ions



(e) Balance charge by adding electrons



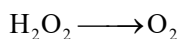
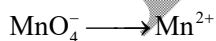
(f) Write the net reaction by adding two half reactions after balancing electrons.



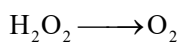
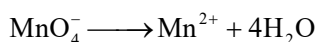
(a) Split the redox reaction



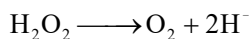
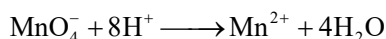
(b) Balance all atoms except H and O



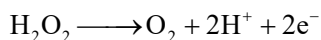
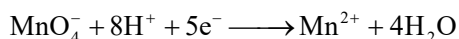
(c) Balance O atoms by adding H<sub>2</sub>O



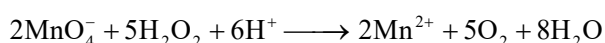
- (d) Balance H atoms by adding H<sup>+</sup> ions



- (e) Balance charge by adding electrons

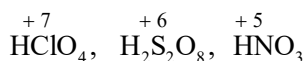


- (f) Write net redox reaction after balancing electrons

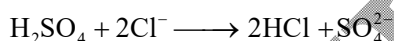


OR

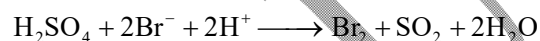
- (i) Species in which the central atom is in the highest oxidation state only act as an oxidising agent.



- (ii) Reaction of H<sub>2</sub>SO<sub>4</sub> with chloride salt is a non-redox reaction and gives pungent fumes of HCl

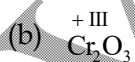


With Br<sup>-</sup>, H<sub>2</sub>SO<sub>4</sub> act as an oxidizing agent and oxidises Br<sup>-</sup> to Br<sub>2</sub> (red vapour) and itself reduced to form SO<sub>2</sub>.

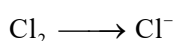
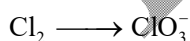


This is redox reaction.

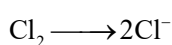
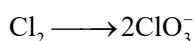
15. (i) (a)  $\overset{+II}{\text{NiSO}_4}$



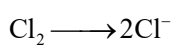
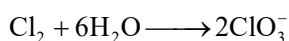
- (ii) This is an example of disproportionation reaction. First of all, we split the redox reaction



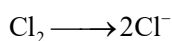
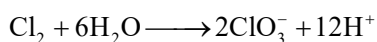
- (a) Balance all atoms except O and H



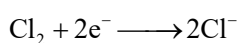
- (b) Balance O atoms by adding H
- <sub>2</sub>
- O



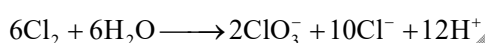
- (c) Balance H atoms by adding H
- <sup>+</sup>
- ions



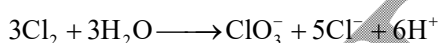
- (d) Balance charge by adding electrons



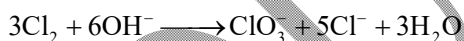
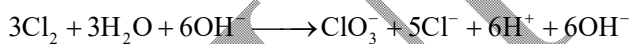
- (e) Add two half reactions after balancing electrons



or



- (f) Since the medium is basic, we add OH
- <sup>-</sup>
- ions equals to H
- <sup>+</sup>
- ions on both sides



OR

- (i) This is an example of disproportionation reaction

Oxidation state of C in (CN)<sub>2</sub> is

$$2x + 2(-3) = 0$$

$$2x - 6 = 0$$

$$x = +3$$

Oxidation state of C in CN<sup>-</sup> is

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

$$x = +2$$

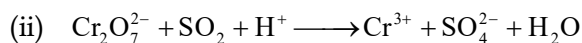
and oxidation state of C in CNO<sup>-</sup> ion is

$$x + 1(-3) + 1(-2) = -1$$

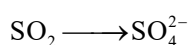
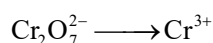
$$x - 3 - 2 = -1$$

$$x = +4$$

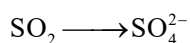
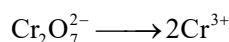
So, oxidation state of carbon is changing from +3 in  $(\text{CN})_2$  to +2 in  $\text{CN}^-$  and +4 in  $\text{CNO}^-$ . Hence, the reaction is a disproportionation reaction.



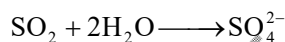
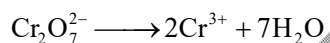
(a) Split the redox reaction



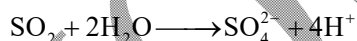
(b) Balance all atoms except H and O



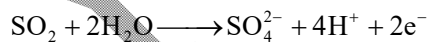
(c) Balance O atoms by adding  $\text{H}_2\text{O}$



(d) Balance H atoms by adding  $\text{H}^+$  ions



(e) Balance charge by adding electrons



(f) Write net redox reaction after balancing electrons.

