

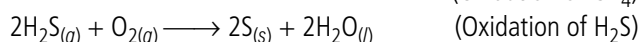
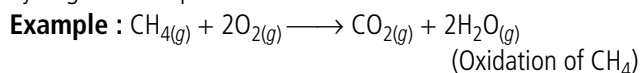
Redox Reactions



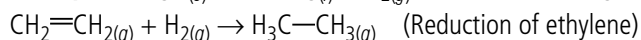
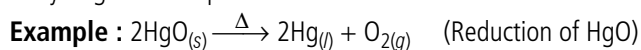
TRY YOURSELF

ANSWERS

1. The term "oxidation" is defined as the addition of oxygen/electronegative element to a substance or removal of hydrogen/electropositive element from a substance.

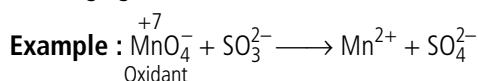


2. The term "reduction" is defined as the removal of oxygen/electronegative element from a substance or addition of hydrogen/electropositive element to a substance.



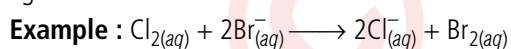
3. BaSO_4 is reduced to BaS because of the removal of oxygen from BaSO_4 .

4. A substance which readily accept electrons to oxidise other substance, as a result itself gets reduced is called oxidising agent or an oxidant.

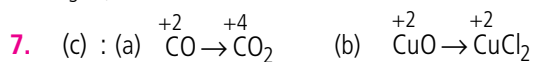
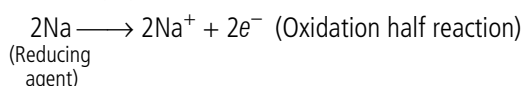
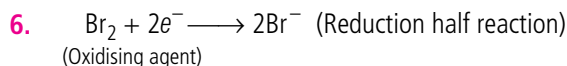


In the above reaction, the Mn reduce from +7 (in MnO_4^-) to +2 (in Mn^{2+}). Therefore, MnO_4^- is the oxidising agent in this reaction.

5. A substance which readily lose electrons to reduce other substance, as a result itself gets oxidised is called reducing agent or reductant.



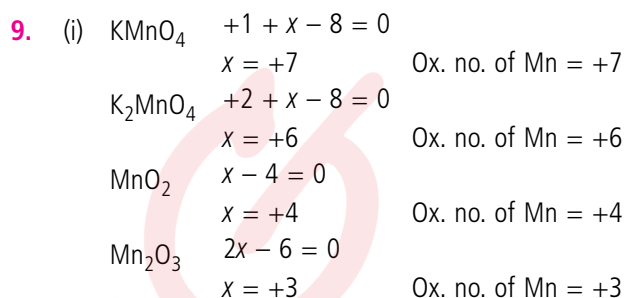
Br^- is reducing agent as in the above reaction oxidation state of Br changes from -1 (in Br^-) to 0 (in Br_2). This means that Br^- oxidises to Br_2 .



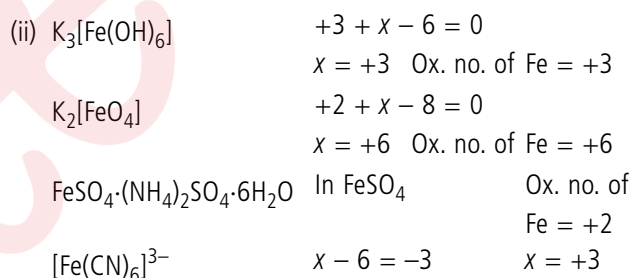
Only in (c), O.N. of hydrogen decreases from + 1 to 0 and hence H_2O gets reduced to H_2 .

8. The decreasing order of electron-releasing tendency or reactivity is $\text{Zn} > \text{Cu} > \text{Ag}$. Such order is found from experiments

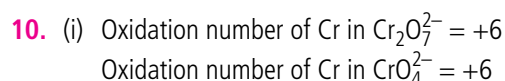
performed by comparing the reactivity of metals and their aqueous solutions. It is found out that Zinc (Zn) reduces both Cu^{2+} to copper (Cu) and $\text{Ag}^+(aq)$ to silver ($\text{Ag}(s)$). While copper only reduces $\text{Ag}^+(aq)$ to silver, $\text{Ag}(s)$.



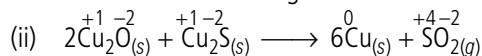
The highest oxidation number for Mn is in KMnO_4 (+7).



The least oxidation number for Fe is in mohl salt $\text{FeSO}_4 \cdot (\text{NH}_4)_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$ *i.e.*, +2.



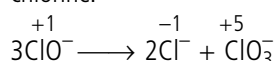
Since, the oxidation number of chromium neither increased nor decreased, therefore, this reaction cannot be regarded as a redox reaction. The oxidation number of other elements, *i.e.*, O and H also remains unchanged.

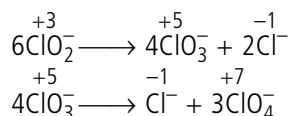


Oxidation state of sulphur changes from -2 in Cu_2S to +4 in SO_2 *i.e.*, it is getting oxidised and hence, Cu_2S acts as a reductant or reducing agent.

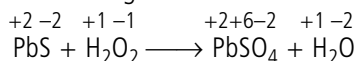
Oxidation state of copper changes from +1 in Cu_2O to zero in elemental copper *i.e.*, It is getting reduced and hence, Cu_2O acts as an oxidant or oxidising agent.

11. Out of the given, ClO_4^- does not undergo disproportionation reaction because in this oxidation state of +7 is highest for chlorine.

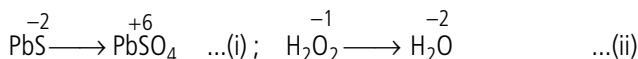




12. Writing oxidation numbers of all the atoms



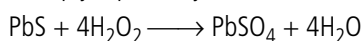
The O.N. of S is increased and that of O is decreased.



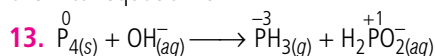
Increase in O.N. of S = 8 units per PbS molecule

Decrease in O.N. of O = 1 unit per 1/2 H₂O₂ molecule
= 2 units per H₂O₂ molecule

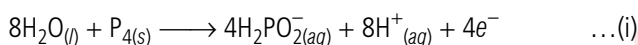
Multiply eqn. (ii) by 4 as to make increase in O.N.



The number of moles of water required in the balanced chemical equation is 4.



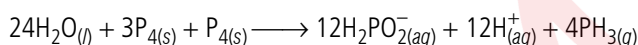
Oxidation half-reaction,



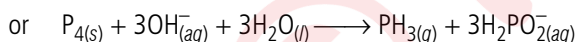
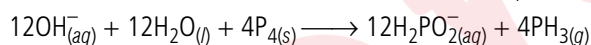
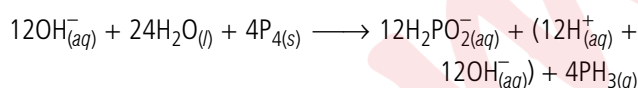
Reduction half-reaction,



On multiplying equation (i) by 3 and adding it to equation (ii), we have

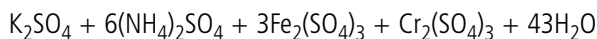
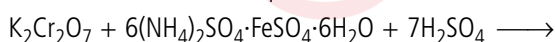


Adding OH⁻ ions on both sides, we have



14. No. of millimoles of Mohr's salt present in 72 mL of 0.5 M solution = 72 × 0.5 = 36

The balanced chemical equation for the redox reaction is



6 moles of Mohr's salt are oxidised by 1 mole of K₂Cr₂O₇

36 millimoles of Mohr's salt are oxidised by $\frac{1}{6} \times 36$

= 6 millimoles of K₂Cr₂O₇

15. Cl₂ is a stronger oxidising agent than I₂, therefore, when Cl₂ is passed through KI solution, Cl₂ gets reduced to colourless Cl⁻ ions while I⁻ ions gets oxidised to violet coloured iodine.



The I₂ produced forms a blue coloured complex with starch which is responsible to impart blue colour to the solution.

16. Meq. of Na₂CO₃ · xH₂O in 20 mL = 19.8 × $\frac{1}{10}$

Meq. of Na₂CO₃ · xH₂O in 100 mL = 19.8 × $\frac{1}{10} \times 5$

$$\therefore \frac{W}{E} \times 1000 = 19.8 \times \frac{1}{10} \times 5$$

$$\frac{0.7}{M/2} \times 1000 = \frac{19.8}{2}$$

$$\therefore M = 141.41$$

$$23 \times 2 + 12 + 3 \times 16 + 18x = 141.41, \quad \therefore x = 2$$

17. KCl cannot be used as an electrolyte in the salt bridge of the given cell because Cl⁻ ion will combine with Ag⁺ ions to form white precipitate of AgCl.

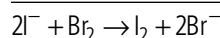
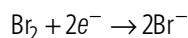
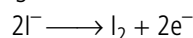
18. Anodic reaction : 2H₂O_(l) → O_{2(g)} + 4H⁺_(aq) + 4e⁻

Cathodic reaction : 2H₂O_(l) + 2e⁻ → H_{2(g)} + 2OH⁻_(aq)

19. E° of H₂O is greater than that of Na⁺ ions, reduction of H₂O takes place preferentially to evolve H₂ gas at the cathode.

20. E° of Br₂ is higher than that of I₂. Therefore, Br₂ has a higher tendency to accept electrons than I₂. Conversely I⁻ ion has a higher tendency to lose electrons than Br⁻ ions.

Therefore, the following reaction will occur



I⁻ ion will get oxidised to I₂ while Br₂ will be reduced to Br⁻ ions.

