

# Redox Reactions

## Multiple Choice Questions (MCQs)

**Q. 1** Which of the following is not an example of redox reaction?

- (a)  $\text{CuO} + \text{H}_2 \longrightarrow \text{Cu} + \text{H}_2\text{O}$   
 (b)  $\text{Fe}_2\text{O}_3 + 3\text{CO} \longrightarrow 2\text{Fe} + 3\text{CO}_2$   
 (c)  $2\text{K} + \text{F}_2 \longrightarrow 2\text{KF}$   
 (d)  $\text{BaCl}_2 + \text{H}_2\text{SO}_4 \longrightarrow \text{BaSO}_4 + 2\text{HCl}$

**Thinking Process**

*Redox reactions represent those reactions which involve change in oxidation number of the interacting species. (i.e., oxidation and reduction)*

**Ans. (d)** Following are the examples of redox reaction

- (a)  $\text{CuO} + \text{H}_2 \longrightarrow \text{Cu} + \text{H}_2\text{O}$   
 (b)  $\text{Fe}_2\text{O}_3 + 3\text{CO} \longrightarrow 2\text{Fe} + 3\text{CO}_2$   
 (c)  $2\text{K} + \text{F}_2 \longrightarrow 2\text{KF}$

Option (d) is not an example of redox reaction.

**Q. 2** The more positive the value of  $E^\ominus$ , 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.

$$E^\ominus \text{ values: } \text{Fe}^{3+}/\text{Fe}^{2+} = +0.77$$

$$\text{I}_2(\text{s})/\text{I}^- = +0.54;$$

$$\text{Cu}^{2+}/\text{Cu} = +0.34; \text{Ag}^+/\text{Ag} = 0.80 \text{ V}$$

- (a)  $\text{Fe}^{3+}$                       (b)  $\text{I}_2(\text{s})$                       (c)  $\text{Cu}^{2+}$                       (d)  $\text{Ag}^+$

**Ans. (d)** Given that,  $E^\ominus$  values of

$$\text{Fe}^{3+}/\text{Fe}^{2+} = +0.77 \text{ V}$$

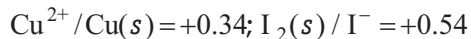
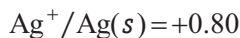
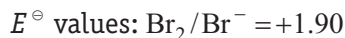
$$\text{I}_2(\text{s})/\text{I}^- = +0.54 \text{ V}$$

$$\text{Cu}^{2+}/\text{Cu} = +0.34 \text{ V}$$

$$\text{Ag}^+/\text{Ag} = +0.80 \text{ V}$$

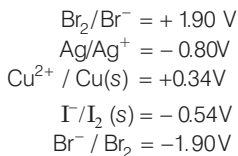
Since,  $E^\ominus$  of the redox couple  $\text{Ag}^+/\text{Ag}$  is the most positive, i.e., 0.80 V, therefore,  $\text{Ag}^+$  is the strongest oxidising agent.

**Q. 3**  $E^\ominus$  values of some redox couples are given below. On the basis of these values choose the correct option.

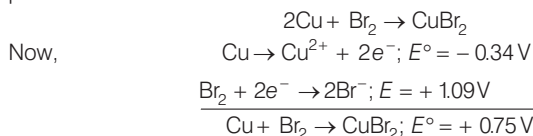


- (a) Cu will reduce  $\text{Br}^-$                                 (b) Cu will reduce Ag  
 (c) Cu will reduce  $\text{I}^-$                                 (d) Cu will reduce  $\text{Br}_2$

**Ans. (d)** Given that  $E^\ominus$  values of



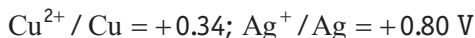
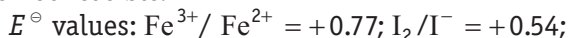
The  $E^\ominus$  values show that copper will reduce  $\text{Br}_2$ , if the  $E^\ominus$  of the following redox reaction is positive.



Since,  $E^\ominus$  of this reaction is positive, therefore, Cu can reduce  $\text{Br}_2$ .

While other reaction will give negative value.

**Q. 4** Using the standard electrode potential, find out the pair between which redox reaction is not feasible.

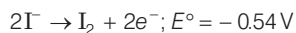
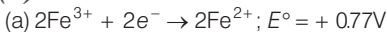


- (a)  $\text{Fe}^{3+}$  and  $\text{I}^-$                                         (b)  $\text{Ag}^+$  and Cu  
 (c)  $\text{Fe}^{3+}$  and Cu                                      (d) Ag and  $\text{Fe}^{3+}$

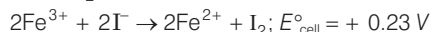
**Thinking Process**

Calculate the  $E^\ominus_{\text{cell}}$  of the four redox reactions. If  $E^\ominus_{\text{cell}}$  of a reaction is negative, that reaction will not occur.

**Ans. (d)**



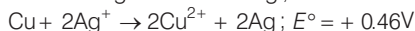
(sign of  $E^\ominus$  is reversed)



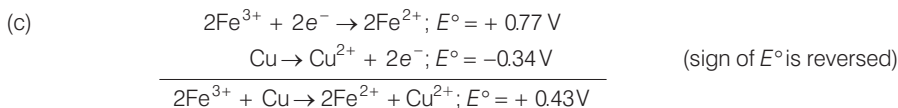
This reaction is feasible since  $E^\ominus_{\text{cell}}$  is positive.



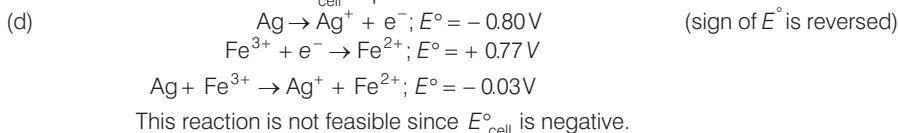
(sign of  $E^\ominus$  has been reversed)



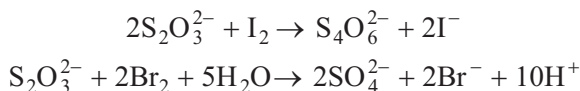
This reaction is feasible since  $E^\ominus_{\text{cell}}$  is positive.



This reaction is feasible since  $E^{\circ}_{\text{cell}}$  is positive.

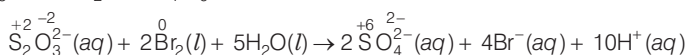
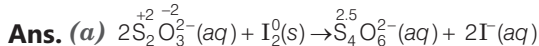


**Q. 5** Thiosulphate reacts differently with iodine and bromine in the reactions given below



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

- Bromine is a stronger oxidant than iodine
- Bromine is a weaker oxidant than iodine
- Thiosulphate undergoes oxidation by bromine and reduction by iodine in these reactions
- Bromine undergoes oxidation and iodine undergoes reduction in these reactions



Bromine being stronger oxidising agent than  $\text{I}_2$ , oxidises S of  $\text{S}_2\text{O}_3^{2-}$  to  $\text{SO}_4^{2-}$  whereas  $\text{I}_2$  oxidises it only into  $\text{S}_4\text{O}_6^{2-}$  ion.

**Q. 6** The oxidation number of an element in a compound is evaluated on the basis of certain rules. Which of the following is incorrect in this respect?

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

**Ans. (a)** Oxidation number of hydrogen is always +1 is a wrong rule since, it is +1 in hydrogen halides, -1 in hydrides and zero in  $\text{H}_2$  molecule.

All the other three statements (b), (c) and (d) are correct.

**Q. 7** In which of the following compounds, an element exhibits two different oxidation states?

- $\text{NH}_2\text{OH}$
- $\text{NH}_4\text{NO}_3$
- $\text{N}_2\text{H}_4$
- $\text{N}_3\text{H}$

**Ans. (b)**  $\text{NH}_4\text{NO}_3$  is actually  $\text{NH}_4^{+}$  and  $\text{NO}_3^{-}$ . It is an ionic compound.

The oxidation number of nitrogen in the two species is different as shown below

Let, oxidation number of N in  $\overset{+}{\text{N}}\text{H}_4$  is  $x$ .

$$\Rightarrow x + (4 \times 1) = +1$$

$$\text{or } x + 4 = +1 \text{ or } x = -3$$

Let, oxidation number of N in  $\text{NO}_3^-$  is  $x$

$$\Rightarrow x + (3 \times -2) = -1 \text{ or } x - 6 = -1 \text{ or } x = +5$$

**Q. 8** Which of the following arrangements represent increasing oxidation number of the central atom ?

- (a)  $\text{CrO}_2^-$ ,  $\text{ClO}_3^-$ ,  $\text{CrO}_4^{2-}$ ,  $\text{MnO}_4^-$       (b)  $\text{ClO}_3^-$ ,  $\text{CrO}_4^{2-}$ ,  $\text{MnO}_4^-$ ,  $\text{CrO}_2^-$   
 (c)  $\text{CrO}_2^-$ ,  $\text{ClO}_3^-$ ,  $\text{MnO}_4^-$ ,  $\text{CrO}_4^{2-}$       (d)  $\text{CrO}_4^{2-}$ ,  $\text{MnO}_4^-$ ,  $\text{CrO}_2^-$ ,  $\text{ClO}_3^-$

**Ans. (a)** Writing the oxidation number (O.N.) of Cr, Cl and Mn on each species in the four set of ions, then,

- (a)  $\overset{+3}{\text{Cr}}\text{O}_2^-$ ,  $\overset{+5}{\text{Cl}}\text{O}_3^-$ ,  $\overset{+6}{\text{Cr}}\text{O}_4^{2-}$ ,  $\overset{+7}{\text{Mn}}\text{O}_4^-$       (b)  $\overset{+5}{\text{Cl}}\text{O}_3^-$ ,  $\overset{+6}{\text{Cr}}\text{O}_4^{2-}$ ,  $\overset{+7}{\text{Mn}}\text{O}_4^-$ ,  $\overset{+3}{\text{Cr}}\text{O}_2^-$   
 (c)  $\overset{+3}{\text{Cr}}\text{O}_2^-$ ,  $\overset{+5}{\text{Cl}}\text{O}_3^-$ ,  $\overset{+7}{\text{Mn}}\text{O}_4^-$ ,  $\overset{+6}{\text{Cr}}\text{O}_4^{2-}$       (d)  $\overset{+6}{\text{Cr}}\text{O}_4^{2-}$ ,  $\overset{+7}{\text{Mn}}\text{O}_4^-$ ,  $\overset{+3}{\text{Cr}}\text{O}_2^-$ ,  $\overset{+5}{\text{Cl}}\text{O}_3^-$

Only in the arrangement (a), the O.N. of central atom increases from left to right, therefore, option (a) is correct.

**Q. 9** 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)  $3d^5 4s^2$

**Ans. (d)** Highest oxidation number of any transition element =  $(n-1)d$  electrons +  $ns$  electrons. Therefore, large the number of electrons in the  $3d$ -orbitals, higher is the maximum 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)  $3d^5 4s^2 = 5 + 2 = 7$

Thus, option (d) is correct.

**Q. 10** Identify disproportionation reaction

- (a)  $\text{CH}_4 + 2\text{O}_2 \longrightarrow \text{CO}_2 + 2\text{H}_2\text{O}$   
 (b)  $\text{CH}_4 + 4\text{Cl}_2 \longrightarrow \text{CCl}_4 + 4\text{HCl}$   
 (c)  $2\text{F}_2 + 2\text{OH}^- \longrightarrow 2\text{F}^- + \text{OF}_2 + \text{H}_2\text{O}$   
 (d)  $2\text{NO}_2 + 2\text{OH}^- \longrightarrow \text{NO}_2^- + \text{NO}_3^- + \text{H}_2\text{O}$

**Ans. (d)** Reactions in which the same substance is oxidised as well as reduced are called disproportionation reactions. Writing the O.N. of each element above its symbol in the given reactions

- (a)  $\overset{-4}{\text{C}}\overset{+1}{\text{H}}_4 + 2\overset{0}{\text{O}}_2 \longrightarrow \overset{+4}{\text{C}}\overset{-2}{\text{O}}_2 + 2\overset{+1}{\text{H}}\overset{-2}{\text{O}}$   
 (b)  $\overset{-4}{\text{C}}\overset{+1}{\text{H}}_4 + 4\overset{0}{\text{Cl}}_2 \longrightarrow \overset{+4}{\text{C}}\overset{-1}{\text{Cl}}_4 + 4\overset{+1}{\text{H}}\overset{-1}{\text{Cl}}$   
 (c)  $2\overset{0}{\text{F}}_2 + 2\overset{-2}{\text{O}}\overset{+1}{\text{H}} \longrightarrow 2\overset{-1}{\text{F}} + \overset{+2}{\text{O}}\overset{-1}{\text{F}} + \overset{+1}{\text{H}}\overset{-2}{\text{O}}$   
 (d)  $2\overset{+4}{\text{N}}\overset{-2}{\text{O}}_2 + 2\overset{-2}{\text{O}}\overset{+1}{\text{H}} \longrightarrow \overset{+3}{\text{N}}\overset{-2}{\text{O}}_2 + \overset{+5}{\text{N}}\overset{-2}{\text{O}}_3 + \overset{+1}{\text{H}}\overset{-2}{\text{O}}$

Thus, in reaction (d), N is both oxidised as well as reduced since the O.N. of N increases from +4 in  $\text{NO}_2$  to +5 in  $\text{NO}_3^-$  and decreases from +4 in  $\text{NO}_2$  to +3 in  $\text{NO}_2^-$ .

**Q. 11** Which of the following elements does not show disproportionation tendency ?

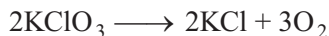
- (a) Cl                      (b) Br                      (c) F                      (d) I

**Ans. (c)** Being the most electronegative element, F can only be reduced and hence it always shows an oxidation number of  $-1$ . Further, due to the absence of  $d$ -orbitals it cannot be oxidised and hence it does not show positive oxidation numbers.

In other words, F cannot be oxidised as well as reduced simultaneously and hence does not show disproportionation reactions.

## Multiple Choice Questions (More Than One Options)

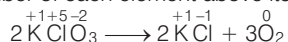
**Q. 12** Which of the following statement(s) is/are **not** true about the following decomposition reaction?



- (a) Potassium is undergoing oxidation  
 (b) Chlorine is undergoing oxidation  
 (c) Oxygen is reduced  
 (d) None of the species are undergoing oxidation or reduction

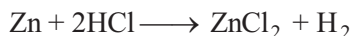
**Ans. (a, b, c, d)**

Write the oxidation number of each element above its symbol, then



- (a) The O.N. of K does not change, K undergoes neither reduction nor oxidation. Thus, option (a) is not correct.  
 (b) The O.N. of chlorine decreases from  $+5$  in  $\text{KClO}_3$  to  $-1$  in  $\text{KCl}$ , hence Cl undergoes reduction.  
 (c) Since, O.N. of oxygen increases from  $-2$  in  $\text{KClO}_3$  to  $0$  in  $\text{O}_2$ , oxygen is oxidised.  
 (d) This statement is not correct because Cl is undergoing reduction and O is undergoing oxidation.

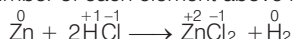
**Q. 13** Identify the correct statement (s) in relation to the following reaction.



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

**Ans. (c, d)**

Writing the oxidation number of each element above its symbol, so that



- (a) The oxidation number of Zn increases from  $0$  in  $\text{Zn}$  to  $+2$  in  $\text{ZnCl}_2$ , therefore, Zn acts as a reductant. Thus, option (a) is incorrect.

- (b) The oxidation number of chlorine does not change, therefore, it neither acts as a reductant nor an oxidant. Therefore, option (b) is incorrect.
- (c) The oxidation number of hydrogen decreases from +1 in  $H^+$  to 0 in  $H_2$ , therefore,  $H^+$  acts as an oxidant. Thus, option (c) is correct.
- (d) As explained in option (a), Zn acts as reductant, therefore, it cannot act as an oxidant. Thus, option (d) is correct.

**Q. 14** The exhibition of various oxidation states by an element is also related to the outer orbital electronic configuration of its atom. Atom(s) having which of the following outermost electronic configurations will exhibit more than one oxidation state in its compounds.

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

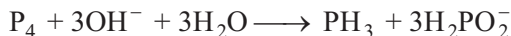
**Ans. (b, c, d)**

**Elements** which have only s-electrons in the valence shell do not show more than one oxidation state. Thus, element with  $3s^1$  as outer electronic configuration shows only one oxidation state of +1.

Transition element such as elements (b), (c) having incompletely filled d-orbitals in the penultimate shell show variable oxidation states. Thus, element with outer electronic configuration as  $3d^1 4s^2$  shows variable oxidation states of +2 and +3 and the element with outer electronic configuration as  $3d^2 4s^2$  shows variable oxidation states of +2, +3 and +4.

**p-Block elements** also show variable oxidation states due to a number of reason such as involvement of d-orbitals and inert pair effect e.g., element (d) with  $3s^2 3p^3$  as (i.e., P) as the outer electronic configuration shows variable oxidation states of +3 and +5 due to involvement of d-orbitals.

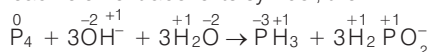
**Q. 15** Identify the correct statements with reference to the given reaction



- (a) Phosphorus is undergoing reduction only  
 (b) Phosphorus is undergoing oxidation only  
 (c) Phosphorus is undergoing oxidation as well as reduction  
 (d) Hydrogen is undergoing neither oxidation nor reduction

**Ans. (c, d)**

Write the O.N. of each element above its symbol, then



In this reaction, O.N. of P increases from 0 in  $P_4$  to +1 in  $H_2PO_2^-$  and decreases to -3 in  $PH_3$ , therefore, P undergoes both oxidation as well as reduction. Thus, options (a) and (b) are wrong and option (c) is correct.

Further, O.N. of H remains +1 in all the compounds, i.e., H neither undergoes oxidation nor reduction. Thus, option (d) is correct.

**Q. 16** Which of the following electrodes will act as anodes, which connected to Standard Hydrogen Electrode ?

- (a)  $Al / Al^{3+}$                        $E^\ominus = -1.66$   
 (b)  $Fe / Fe^{2+}$                        $E^\ominus = -0.44$   
 (c)  $Cu / Cu^{2+}$                        $E^\ominus = +0.34$   
 (d)  $F_2 (g) / 2F^- (aq)$                        $E^\ominus = 0.287$

**Ans. (a, b)**

All electrodes which have negative electrode potentials are stronger reducing agents than  $H_2$  gas and hence acts as anodes when connected to standard hydrogen electrode. Thus,  $Al^{3+} / Al$  ( $E^\ominus = -1.66$  V) and  $Fe^{2+} / Fe$  ( $E^\ominus = -0.44$  V) act as anode.

## Short Answer Type Questions

**Q. 17** The reaction  $Cl_2(g) + 2OH^-(aq) \rightarrow ClO^-(aq) + Cl^-(aq) + H_2O(l)$  represents the process of bleaching. Identify and name the species that bleaches the substances due to its oxidising action.

### Thinking Process

Write the oxidation number of each element above its symbol. and then identify the bleaching reagent by observing the change in oxidation number.

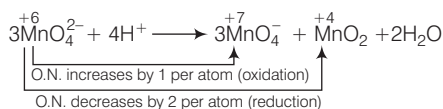
**Ans.**  $Cl_2^0(g) + 2OH^{+1-2}(aq) \rightarrow ClO^{-1-2}(aq) + Cl^{-1}(aq) + H_2O^{+1-2}(l)$

In this reaction, O.N. of Cl increases from 0 (in  $Cl_2$ ) to 1 (in  $ClO^-$ ) as well as decreases from 0 (in  $Cl_2$ ) to -1 (in  $Cl^-$ ). So, it acts both reducing as well as oxidising agent. This is an example of disproportionation reaction. In this reaction,  $ClO^-$  species bleaches the substances due to its oxidising action. [In hypochlorite ion ( $ClO^-$ ) Cl can decrease its oxidation number from +1 to 0 or -1.]

**Note** Disproportionation reactions are a special type of redox reactions. In which an element in one oxidation state is simultaneously oxidised and reduced.

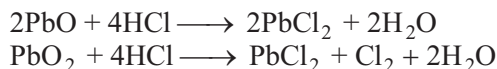
**Q. 18**  $MnO_4^{2-}$  undergoes disproportionation reaction in acidic medium but  $MnO_4^-$  does not. Give reason.

**Ans.** In  $MnO_4^{2-}$ , the oxidation number of Mn is +6. It can increase its oxidation number (to + 7) or decrease its oxidation number (to + 4, + 3, + 2, 0). Hence, it undergoes disproportionation reaction in acidic medium.



In  $MnO_4^-$ , Mn is in its highest oxidation state, i.e., +7. It can only decrease its oxidation number. Hence, it cannot undergo disproportionation reaction.

**Q. 19**  $PbO$  and  $PbO_2$  react with  $HCl$  according to following chemical equations

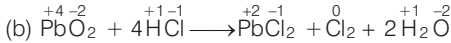


Why do these compounds differ in their reactivity ?

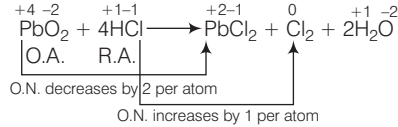
**Ans.** Writing the oxidation number of each element above its symbol in the following reactions



In this reaction, oxidation number of each element remains same hence, it is not a redox reaction. In fact, it is an example of **acid-base reaction**.

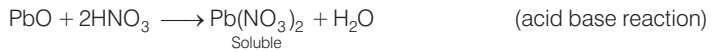


In  $\text{PbO}_2$ , Pb is in +4 oxidation state. Due to inert pair effect Pb in +2 oxidation state is more stable. So, Pb in +4 oxidation state ( $\text{PbO}_2$ ) acts as an oxidising agent. It oxidises  $\text{Cl}^-$  to  $\text{Cl}_2$  and itself gets reduced to  $\text{Pb}^{2+}$ .



**Q. 20** Nitric acid is an oxidising agent and reacts with PbO but it does not react with  $\text{PbO}_2$ . Explain why ?

**Ans.** PbO is a base. It reacts with nitric acid and forms soluble lead nitrate.

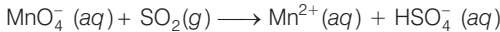


Nitric acid does not react with  $\text{PbO}_2$ . Both of them are strong oxidising agents. In  $\text{HNO}_3$ , nitrogen is in its maximum oxidation state (+5) and in  $\text{PbO}_2$ , lead is in its maximum oxidation state (+4). Therefore, no reaction takes place.

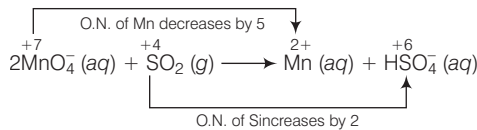
**Q. 21** Write balanced chemical equation for the following reactions.

- (a) Permanganate ion ( $\text{MnO}_4^-$ ) reacts with sulphur dioxide gas in acidic medium to produce  $\text{Mn}^{2+}$  and hydrogen sulphate ion.  
(Balance by ion electron method)
- (b) Reaction of liquid hydrazine ( $\text{N}_2\text{H}_4$ ) with chlorate ion ( $\text{ClO}_3^-$ ) in basic medium produces nitric oxide gas and chloride ion in gaseous state.  
(Balance by oxidation number method)
- (c) Dichlorine heptaoxide ( $\text{Cl}_2\text{O}_7$ ) in gaseous state combines with an aqueous solution of hydrogen peroxide in acidic medium to give chlorite ion ( $\text{ClO}_2^-$ ) and oxygen gas.  
(Balance by ion electron method)

**Ans.** (a) **Ion electron method** Write the skeleton equation for the given reaction.



Find out the elements which undergo change in O.N.

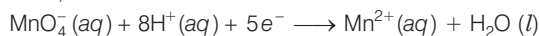


Divide the given skeleton into two half equations.



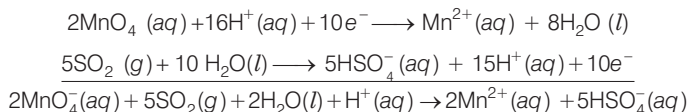
To balance reduction half equation

In acidic medium, balance H and O-atoms

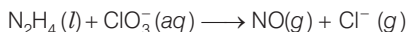




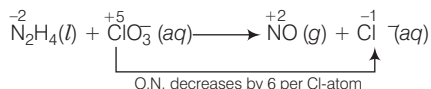
To balance the complete reaction



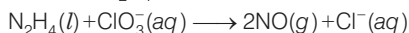
(b) **Oxidation number method** Write the skeleton equation for the given reaction.



O.N. increases by 4 per N-atom



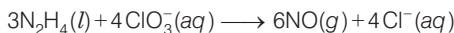
Multiply NO by 2 because in  $\text{N}_2\text{H}_4$  there are 2N atoms



Total increase in O.N. of N =  $2 \times 4 = 8$  ( $8e^-$  lost)

Total decrease in O.N. of Cl =  $1 \times 6 = 6$  ( $6e^-$  gain)

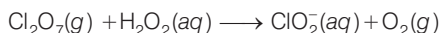
Therefore, to balance increase or decrease in O.N. multiply  $\text{N}_2\text{H}_4$  by 3, 2NO by 3 and  $\text{ClO}_3^-$ ,  $\text{Cl}^-$  by 4



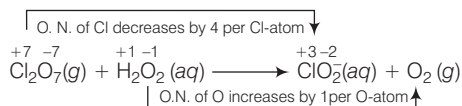
Balance O and H-atoms by adding  $6\text{H}_2\text{O}$  to RHS



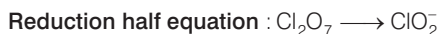
(c) **Ion electron method** Write the skeleton equation for the given reaction.



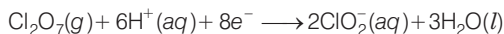
Find out the elements which undergo a change in O.N.



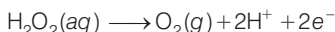
Divide the given skeleton equation into two half equations.



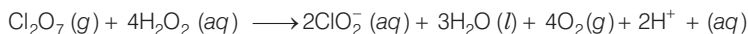
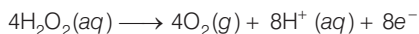
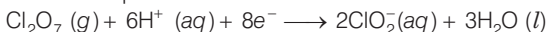
To balance the reduction half equation



To balance the oxidation half equation

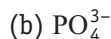


To balance the complete reaction



This represents the balanced redox reaction.

**Q. 22** Calculate the oxidation number of phosphorus in the following species.



**Ans. (a)** Suppose that the O.N. of P in  $\text{HPO}_3^{2-}$  be  $x$ .

$$\text{Then, } 1 + x + 3(-2) = -2$$

$$\text{or, } x + 1 - 6 = -2$$

$$\text{or, } x = +3$$

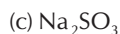
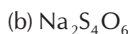
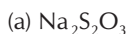
**(b)** Suppose that the O.N. of P in  $\text{PO}_4^{3-}$  be  $x$ .

$$\text{Then, } x + 4(-2) = -3$$

$$\text{or, } x - 8 = -3$$

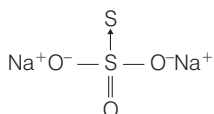
$$\text{or, } x = +5$$

**Q. 23** Calculate the oxidation number of each sulphur atom in the following compounds.



**Ans.** The oxidation number of each sulphur atom in the following compounds are given below

(a)  $\text{Na}_2\text{S}_2\text{O}_3$  Let us consider the structure of  $\text{Na}_2\text{S}_2\text{O}_3$ .



There is a coordinate bond between two sulphur atoms. The oxidation number of acceptor S-atom is  $-2$ . Let, the oxidation number of other S-atom be  $x$ .

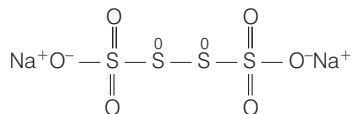
$$2(+1) + 3 \times (-2) + x + 1(-2) = 0$$

For Na      For O-atoms                      For coordinate S-atom

$$x = +6$$

Therefore, the two sulphur atoms in  $\text{Na}_2\text{S}_2\text{O}_3$  have  $-2$  and  $+6$  oxidation number.

(b)  $\text{Na}_2\text{S}_4\text{O}_6$  Let us consider the structure of  $\text{Na}_2\text{S}_4\text{O}_6$ .



In this structure, two central sulphur atoms have zero oxidation number because electron pair forming the S—S bond remain in the centre. Let, the oxidation number of (remaining S-atoms) S-atom be  $x$ .

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

For Na      For O

$$2 - 12 + 2x = 0 \text{ or } x = +\frac{10}{2} = +5$$

Therefore, the two central S-atoms have zero oxidation state and two terminal S-atoms have  $+5$  oxidation state each.

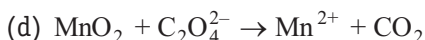
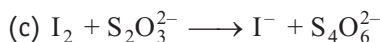
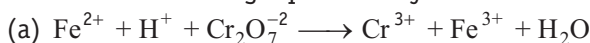
(c)  $\text{Na}_2\text{SO}_3$  Let the oxidation number of S in  $\text{Na}_2\text{SO}_3$  be  $x$ .

$$2(+1) + x + 3(-2) = 0 \text{ or } x = +4$$

(d)  $\text{Na}_2\text{SO}_4$  Let the oxidation number of S be  $x$ .

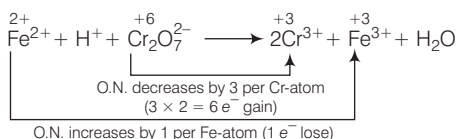
$$2(+1) + x + 4(-2) = 0 \text{ or } x = +6$$

**Q. 24** Balance the following equations by the oxidation number method.



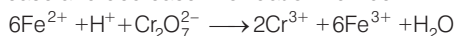
**Ans.** Oxidation number method

(a)

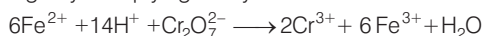


(Multiply  $\text{Cr}^{3+}$  by 2 because there are 2Cr atoms in  $\text{Cr}_2\text{O}_7^{2-}$  ion.)

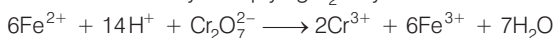
Balance increase and decrease in oxidation number.



Balance charge by multiplying  $\text{H}^+$  by 14.

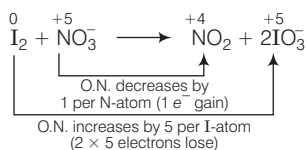


Balance H and O-atoms by multiplying  $\text{H}_2\text{O}$  by 7.

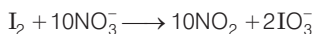


This represents a balanced redox reaction.

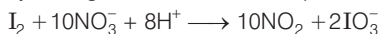
(b)



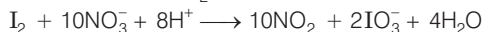
Balance increase and decrease in oxidation number



Balance charge by writing  $8\text{H}^+$  in LHS of the equation.



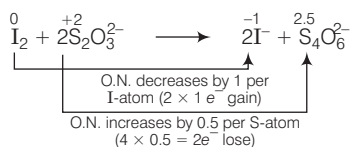
Balance H-atoms by writing  $4\text{H}_2\text{O}$  in RHS of the equation.



Oxygen atoms are automatically balanced.

This represents a balanced redox reaction.

(c)

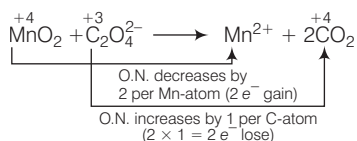


(Multiply  $\text{S}_2\text{O}_3^{2-}$  by 2 because there are 4 S-atoms in  $\text{S}_4\text{O}_6^{2-}$  ion.)

Increase and decrease in oxidation number is already balanced. Charge and oxygen atoms are also balanced.

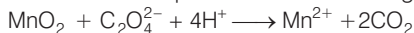
This represents a balanced redox reaction.

(d)

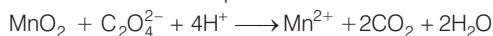


Increase and decrease in oxidation number is already balanced.

Add  $4\text{H}^+$  towards LHS of the equation to balance charge.

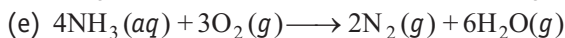
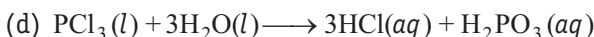
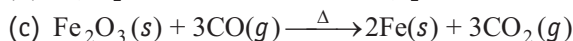
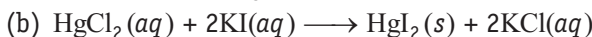
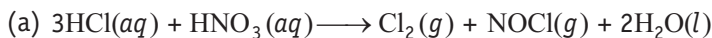


Add  $2\text{H}_2\text{O}$  towards RHS of the equation to balance H-atoms

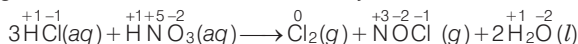


This represents a balanced redox reaction.

**Q. 25** Identify the redox reaction out of the following reactions and identify the oxidising and reducing agents in them.



**Ans. (a)** Writing the O.N. on each atom above its symbol, then

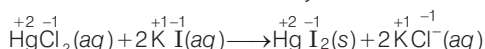


Here, the O.N. of Cl increases from  $-1$  in HCl to  $0$  in  $\text{Cl}_2$ , therefore,  $\text{Cl}^-$  is oxidised and hence HCl acts as the reducing agent.

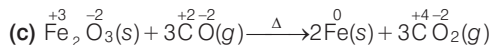
The O.N. of N decreases from  $+5$  in  $\text{HNO}_3$  to  $+3$  in  $\text{NOCl}$ , therefore,  $\text{HNO}_3$  acts as the oxidising agent.

Thus, this reaction is a redox reaction.

(b) Writing the O.N. of each atom above its symbol, we have,



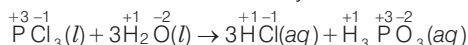
Here, the O.N. of none of the atoms undergo a change, therefore, this reaction is not a redox reaction.



Here, O.N. of Fe decreases from  $+3$  in  $\text{Fe}_2\text{O}_3$  to  $0$  in Fe, therefore,  $\text{Fe}_2\text{O}_3$  acts as an oxidising agent. Further, O.N. of C increases from  $+2$  in CO to  $+4$  in  $\text{CO}_2$ , therefore, CO acts as a reducing agent.

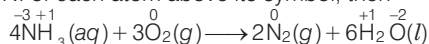
Thus, this reaction is an example of redox reaction.

(d) Writing the O.N. of each atom above its symbol, then



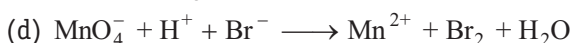
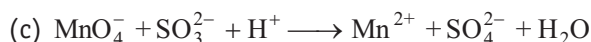
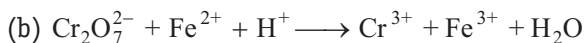
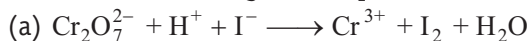
Here, O.N. of none of the atoms undergo a change, therefore, this reaction is not a redox reaction.

(e) Writing the O.N. of each atom above its symbol, then



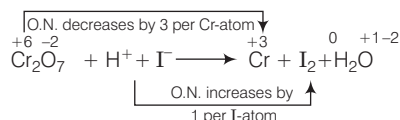
Here, O.N. of N increases from  $-3$  to  $0$  in  $N_2$ , therefore,  $NH_3$  acts as a reducing agent. Further, O.N. of O decreases from  $0$  in  $O_2$  to  $-2$  in  $H_2O$ , therefore,  $O_2$  acts as an oxidising agent. Thus, this reaction is a redox reaction.

**Q. 26** Balance the following ionic equations.



**Ans. (a)** Write the O. N. of all atoms above their respective symbols.

O. N. decreases by, 3 per Cr-atom

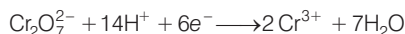


Divide the given equation into two half reactions

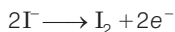
**Reduction half reaction** :  $Cr_2O_7^{2-} \rightarrow Cr^{3+}$

**Oxidation half reaction** :  $I^- \rightarrow I_2$

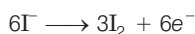
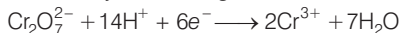
To balance reduction half reaction.



To balance oxidation half reaction

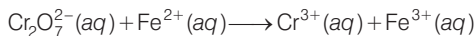


To balance the reaction by electrons gained and lost

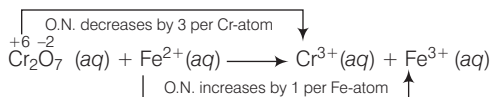


This gives the final balanced ionic equations.

**(b)** Write the skeletal equation of the given reaction



Write the O. N. of all the elements above their respective symbols.

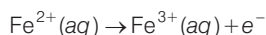


Divide the given equation into two half reactions

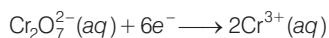
**Oxidation half reaction** :  $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq)$

**reduction half reaction** :  $Cr_2O_7^{2-}(aq) \rightarrow Cr^{3+}(aq)$

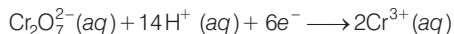
To balance oxidation half reaction



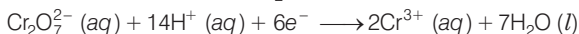
To balance reduction half reaction



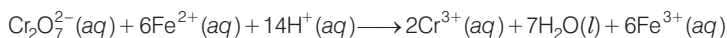
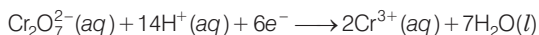
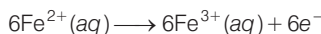
Balance charge by adding  $H^+$  ions.



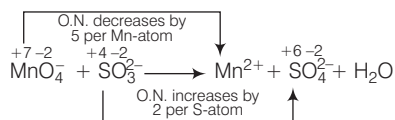
Balance O atoms by adding  $H_2O$  molecules



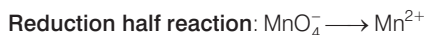
To balance the reaction



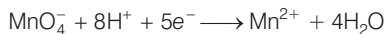
(c) Write the O. N. of all atoms above their respective symbols.



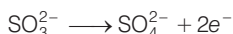
Divide the skeleton equation into two half-reactions.



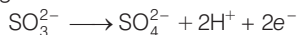
To balance reduction half reaction



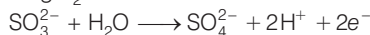
To balance oxidation half reaction



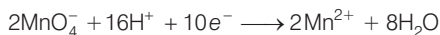
Balance charge by adding  $H^+$  ions.



Balance O-atoms by adding  $H_2O$  molecules

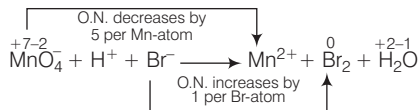


To balance the reaction

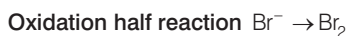


This represents the correct balanced redox equation.

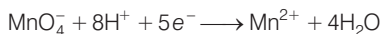
(d) Write the O. N. of all the atoms above their respective symbols.



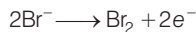
Divide skeleton equation into two half reactions



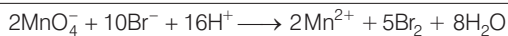
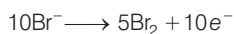
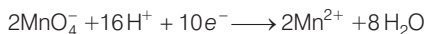
To balance reduction half reaction



To balance oxidation half reaction



To balance the reaction



This represents the correct balanced ionic equation.

## Matching The Columns

**Q. 27** Match Column I with Column II for the oxidation states of the central atoms.

	Column I	Column II
A.	$\text{Cr}_2\text{O}_7^{2-}$	1. +3
B.	$\text{MnO}_4^-$	2. +4
C.	$\text{VO}_3^-$	3. +5
D.	$\text{FeF}_6^{3-}$	5. +6
		6. +7

**Ans.** A. → (4)    B. → (5)    C. → (3)    D. → (1)

Suppose that  $x$  be the oxidation states of central atoms.

A. Oxidation number of Cr in  $\text{Cr}_2\text{O}_7^{2-}$

$$2x + 7(-2) = -2$$

$$2x - 14 = -2$$

$$2x = +12$$

$$x = +6$$

B. Oxidation number of Mn in  $\text{MnO}_4^-$

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

$$x - 8 = -1$$

$$x = +7$$

C. Oxidation number of V in  $\text{VO}_3^-$

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

$$x - 6 = -1$$

$$x = +5$$

D. Oxidation number of Fe in  $\text{FeF}_6^{3-}$

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

$$x - 6 = -3$$

or

$$x = +3$$

**Q. 28** Match the items in Column I with relevant items in Column II.

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 ( $H^+$ )	3. +1
D. Oxidation number of fluorine in NaF	4. 0
E. Ions having negative charge	5. Cation
	6. Anion

**Ans.** A. → (5)    B. → (4)    C. → (3)    D. → (2)    E. → (6)

- A. Ions having positive charge — Cation  
 B. The sum of oxidation number of all atoms in a neutral molecule — Zero  
 C. Oxidation number of hydrogen ion ( $H^+$ ) — +1  
 D. Oxidation number of fluorine in NaF — -1  
 E. Ions having negative charge — Anion

## Assertion and Reason

*In the following questions a statement of assertion (A) followed by a statement of reason (R) is given. Choose the correct option out of the choices given below in each question.*

**Q. 29 Assertion (A)** Among halogens fluorine is the best oxidant.

**Reason (R)** Fluorine is the most electronegative atom.

- (a) Both A and R are true and R is the correct explanation of A  
 (b) Both A and R are true but R is not the correct explanation of A  
 (c) A is true but R is false  
 (d) Both A and R are false

**Ans. (b)** Both assertion and reason are true but reason is not the correct explanation of assertion. Among halogen  $F_2$  is the best oxidant because it has the highest  $E^\circ$  value.

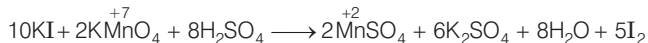
**Q. 30 Assertion (A)** In the reaction between potassium permanganate and potassium iodide, permanganate ions act as oxidising agent.

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

- (a) Both A and R are true and R is the correct explanation of A  
 (b) Both A and R are true but R is not the correct explanation of A  
 (c) A is true but R is false  
 (d) Both A and R are false



**Ans. (c)** Assertion is true but reason is false.



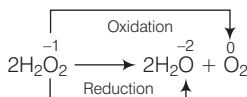
Oxidation state of Mn decreases from +7 to +2.

**Q. 31 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  $\text{O}_2$  and  $-2$  oxidation state in  $\text{H}_2\text{O}$ .

- (a) Both A and R are true and R is the correct explanation of A  
 (b) Both A and R are true but R is not the correct explanation of A  
 (c) A is true but R is false  
 (d) Both A and R are false

**Ans. (a)** Both assertion and reason are true and reason is the correct explanation of assertion.



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

**Q. 32 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  $E_{\text{Fe}^{3+}/\text{Fe}^{2+}}^\ominus$  and  $E_{\text{Cu}^{2+}/\text{Cu}}^\ominus$ ,  $\text{Fe}^{3+}/\text{Fe}^{2+}$  and  $\text{Cu}^{2+}/\text{Cu}$  are redox couples.

- (a) Both A and R are true and R is the correct explanation of A  
 (b) Both A and R are true but R is not the correct explanation of A  
 (c) A is true but R is false  
 (d) Both A and R are false

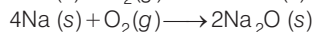
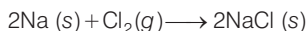
**Ans. (a)** Both assertion and reason are true reason is the correct explanation of assertion.

Redox couple is the combination of oxidised and reduced form of substance. In the representation  $E_{\text{Fe}^{3+}/\text{Fe}^{2+}}^\ominus$  and  $E_{\text{Cu}^{2+}/\text{Cu}}^\ominus$ ,  $\text{Fe}^{3+}/\text{Fe}^{2+}$  and  $\text{Cu}^{2+}/\text{Cu}$  are redox couples.

## Long Answer Type Questions

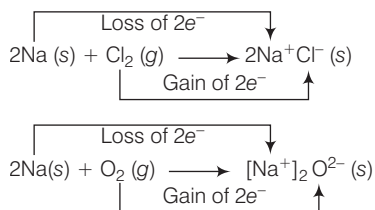
**Q. 33** Explain redox reaction on the basis of electron transfer. Given suitable examples.

**Ans.** As we know that, the reactions

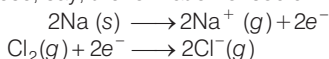


are redox reactions because in each of these reactions sodium is oxidised due to the addition of either oxygen or more electronegative element to sodium. Simultaneously, chlorine and oxygen are reduced because of each of these, the electropositive element sodium has been added.

From our knowledge of chemical bonding we also know that, sodium chloride and sodium oxide are ionic compounds and perhaps better written as  $\text{Na}^+\text{Cl}^-(\text{s})$  and  $(\text{Na}^+)_2\text{O}^{2-}(\text{s})$ . Development of charges on the species produced suggests us to rewrite the above reaction in the following manner



For convenience, each of the above processes can be considered as two separate steps, one involving the loss of electrons and other the gain of electrons. As an illustration, we may further elaborate one of these, say, the formation of sodium chloride.



Each of the above steps is called a half reaction, which explicitly shows involvement of electrons. Sum of the half reactions gives the overall reaction:



The given reactions suggest that half reactions that involved loss of electrons are oxidation reactions. Similarly, the half reactions that involve gain of electrons are called reduction reactions.

It may not be out of context to mention here that the new way of defining oxidation and reduction has been achieved only by establishing a correlation between the behaviour of species as per the classical idea and their interplay in electron-transfer change.

In the given reactions, sodium, which is oxidised, acts as a reducing agent because it donates electron to each of the elements interacting with it and thus helps in reducing them. Chlorine and oxygen are reduced and act as oxidising agents because these accept electrons from sodium.

To summarise, we may mention that

**Oxidation** Loss of electron(s) by any species.

**Reduction** Gain of electron(s) by any species.

**Oxidising agent** Acceptor of electron(s).

**Reducing agent** Donor of electron(s).

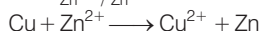
**Q. 34** On the basis of standard electrode potential values, suggest which of the following reactions would take place? (Consult the book for  $E^\ominus$  value)

- $\text{Cu} + \text{Zn}^{2+} \longrightarrow \text{Cu}^{2+} + \text{Zn}$
- $\text{Mg} + \text{Fe}^{2+} \longrightarrow \text{Mg}^{2+} + \text{Fe}$
- $\text{Br}_2 + 2\text{Cl}^- \longrightarrow \text{Cl}_2 + 2\text{Br}^-$
- $\text{Fe} + \text{Cd}^{2+} \longrightarrow \text{Cd} + \text{Fe}^{2+}$

**Ans.** As we know that,

$$\begin{array}{l}
 E^\ominus_{\text{Cu}^{2+}/\text{Cu}} = 0.34 \text{ V}, E^\ominus_{\text{Zn}^{2+}/\text{Zn}} = -0.76 \text{ V}, \\
 E^\ominus_{\text{Mg}^{2+}/\text{Mg}} = -2.37 \text{ V}, E^\ominus_{\text{Fe}^{2+}/\text{Fe}} = -0.74 \text{ V}, \\
 E^\ominus_{\text{Br}_2/\text{Br}^-} = +1.08 \text{ V}, E^\ominus_{\text{Cl}_2/\text{Cl}^-} = +1.36 \text{ V} \\
 E^\ominus_{\text{Cd}^{2+}/\text{Cd}} = -0.44 \text{ V}
 \end{array}$$

(a)  $E^\circ_{\text{Cu}^{2+}/\text{Cu}} = +0.34 \text{ V}$  and  $E^\circ_{\text{Zn}^{2+}/\text{Zn}} = -0.76 \text{ V}$

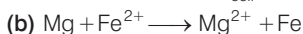


In the given cell reaction, Cu is oxidised to  $\text{Cu}^{2+}$ , therefore,  $\text{Cu}^{2+}/\text{Cu}$  couple acts as anode and  $\text{Zn}^{2+}$  is reduced to Zn, therefore,  $\text{Zn}^{2+}/\text{Zn}$  couple acts as cathode.

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$E^\circ_{\text{cell}} = -0.76 - (+0.34) = -1.10 \text{ V}$$

Negative value of  $E^\circ_{\text{cell}}$  indicates that the reaction will not occur.



$$E^\circ_{\text{Mg}^{2+}/\text{Mg}} = -2.37 \text{ V} \text{ and } E^\circ_{\text{Fe}^{2+}/\text{Fe}} = -0.74 \text{ V}$$

In the given cell reaction, Mg is oxidised to  $\text{Mg}^{2+}$  hence,  $\text{Mg}^{2+}/\text{Mg}$  couple acts as anode and  $\text{Fe}^{2+}$  is reduced to Fe hence,  $\text{Fe}^{2+}/\text{Fe}$  couple acts as cathode.

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$E^\circ_{\text{cell}} = -0.74 - (-2.37) = +1.63 \text{ V}$$

Positive value of  $E^\circ_{\text{cell}}$  indicates that the reaction will occur.



$$E^\circ_{\text{Br}^-/\text{Br}_2} = +1.08 \text{ V} \text{ and } E^\circ_{\text{Cl}^-/\text{Cl}_2} = +1.36 \text{ V}$$

In the given cell reaction,  $\text{Cl}^-$  is oxidised to  $\text{Cl}_2$  hence,  $\text{Cl}^-/\text{Cl}_2$  couple acts as anode and  $\text{Br}_2$  is reduced to  $\text{Br}^-$  hence;  $\text{Br}^-/\text{Br}_2$  couple acts as cathode.

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$E^\circ_{\text{cell}} = +1.08 - (+1.36) = -0.28 \text{ V}$$

Negative value of  $E^\circ_{\text{cell}}$  indicates that the reaction will occur.



$$E^\circ_{\text{Fe}^{2+}/\text{Fe}} = -0.74 \text{ V} \text{ and } E^\circ_{\text{Cd}^{2+}/\text{Cd}} = -0.44 \text{ V}$$

In the given cell reaction, Fe is oxidised to  $\text{Fe}^{2+}$  hence,  $\text{Fe}^{2+}/\text{Fe}$  couple acts as anode and  $\text{Cd}^{2+}$  is reduced to Cd hence,  $\text{Cd}^{2+}/\text{Cd}$  couple acts as cathode.

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$E^\circ_{\text{cell}} = -0.44 - (-0.74) = +0.30 \text{ V}.$$

Positive value  $E^\circ_{\text{cell}}$  indicates that the reaction will occur.

### Q. 35 Why does fluorine not show disproportionation reaction?

**Ans.** In a disproportionation reaction, the same species is simultaneously oxidised as well as reduced. Therefore, for such a redox reaction to occur, the reacting species must contain an element which has atleast three oxidation states.

The element, in reacting species, is present in an intermediate state while lower and higher oxidation states are available for reduction and oxidation to occur (respectively).

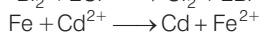
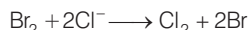
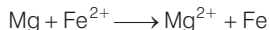
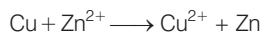
Fluorine is the strongest oxidising agent. It does not show positive oxidation state. That's why fluorine does not show disproportionation reaction.

**Q. 36** Write redox couples involved in the reactions (a) to (d) given in question 34.

**Thinking Process**

*A redox couple represents the oxidised and reduced forms of a substance together taking part in an oxidation or reduction half reaction.*

**Ans.** Given that,



(a)  $\text{Cu}^{2+}/\text{Cu}$  and  $\text{Zn}^{2+}/\text{Zn}$

(b)  $\text{Mg}^{2+}/\text{Mg}$  and  $\text{Fe}^{2+}/\text{Fe}$

(c)  $\text{Br}_2/\text{Br}^-$  and  $\text{Cl}_2/\text{Cl}^-$

(d)  $\text{Fe}^{2+}/\text{Fe}$  and  $\text{Cd}^{2+}/\text{Cd}$

**Q. 37** Find out the oxidation number of chlorine in the following compounds and arrange them in increasing order of oxidation number of chlorine.

$\text{NaClO}_4$ ,  $\text{NaClO}_3$ ,  $\text{NaClO}$ ,  $\text{KClO}_2$ ,  $\text{Cl}_2\text{O}_7$ ,  $\text{ClO}_3$ ,  $\text{Cl}_2\text{O}$ ,  $\text{NaCl}$ ,  $\text{Cl}_2$ ,  $\text{ClO}_2$ .

Which oxidation state is not present in any of the above compounds?

**Ans.** Suppose that the oxidation number of chlorine in these compounds be  $x$ .

$$\text{O.N. of Cl in } \text{NaClO}_4 \therefore +1 + x + 4(-2) = 0 \text{ or, } x = +7$$

$$\text{O.N. of Cl in } \text{NaClO}_3 \therefore +1 + x + 3(-2) = 0 \text{ or, } x = +5$$

$$\text{O.N. of Cl in } \text{NaClO} \therefore +1 + x + 1(-2) = 0 \text{ or, } x = +1$$

$$\text{O.N. of Cl in } \text{KClO}_2 \therefore +1 + x + 2(-2) = 0 \text{ or, } x = +3$$

$$\text{O.N. of Cl in } \text{Cl}_2\text{O}_7 \therefore +2x + 7(-2) = 0 \text{ or, } x = +7$$

$$\text{O.N. of Cl in } \text{ClO}_3 \therefore x + 3(-2) = 0 \text{ or, } x = +6$$

$$\text{O.N. of Cl in } \text{Cl}_2\text{O} \therefore 2x + 1(-2) = 0 \text{ or, } x = +1$$

$$\text{O.N. of Cl in } \text{NaCl} \therefore +1 + x = 0 \text{ or, } x = -1$$

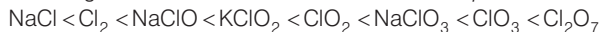
$$\text{O.N. of Cl in } \text{Cl}_2 \therefore 2x = 0 \text{ or, } x = 0$$

$$\text{O.N. of Cl in } \text{ClO}_2 \therefore x + 2(-2) = 0 \text{ or, } x = +4$$

None of these compounds have an oxidation number of +2.

Increasing order of oxidation number of chlorine is :  $-1, 0, +1, +3, +4, +5, +6, +7$

Therefore, the increasing order of oxidation number of Cl in compounds is

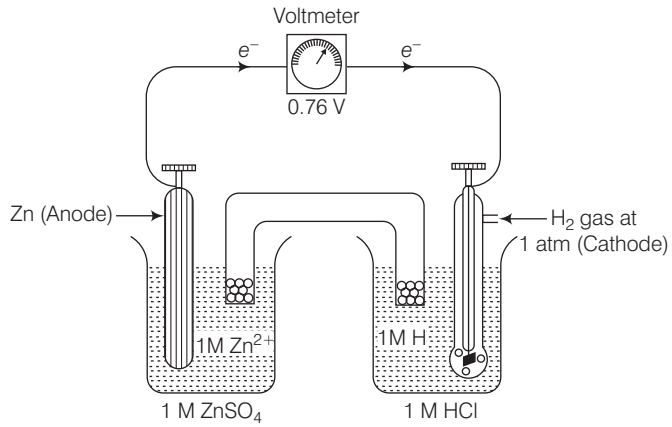


**Q. 38** Which method can be used to find out strength of reductant/oxidant in a solution? Explain with an example.

**Ans.** Measure the electrode potential of the given species by connecting the redox couple of the given species with standard hydrogen electrode. If it is positive, the electrode of the given species acts as reductant and if it is negative, it acts as an oxidant.

Find the electrode potentials of the other given species in the same way, compare the values and determine their comparative strength as a reductant or oxidant.

e.g., measurement of standard electrode potential of  $\text{Zn}^{2+}/\text{Zn}$  electrode using SHE as a reference electrode.



The EMF of the cell comes out to be 0.76 V. (reading of voltmeter is 0.76 V).  $\text{Zn}^{2+}/\text{Zn}$  couple acts as anode and SHE acts as cathode.

$$\begin{aligned} \therefore E_{\text{cell}}^{\circ} &= 0.76 = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} \\ 0.76 &= 0 - E_{\text{anode}}^{\circ} \\ E_{\text{anode}}^{\circ} &= -0.76 \text{ V} \\ E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} &= -0.76 \text{ V}. \end{aligned}$$