

SUBJECT : CHEMISTRY

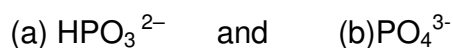
CHAPTER-08 : REDOX REACTIONS

QUESTIONS CARRYING ONE MARK:

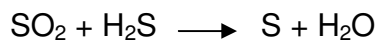
1. Define 'oxidation' in terms of electron transfer.
2. Give the electronic interpretation of 'reduction'.
3. What is an oxidizing agent (or oxidant)?
4. Which is the most powerful oxidizing agent?
5. What is a reducing agent (or reductant)?
6. Which is the most powerful reducing agent?
7. Complete the following equation: $2\text{Fe}^{2+} + 2\text{H}^+ + \text{H}_2\text{O}_2 \longrightarrow \text{.....} + 2\text{H}_2\text{O}$.
8. Define oxidation number. (or oxidation state).
9. Calculate the oxidation number of Cr in $\text{Cr}_2\text{O}_7^{2-}$.
10. Calculate the oxidation number of Mn in KMnO_4 .
11. What is the oxidation number (or oxidation state) of an element?
12. What happens to the oxidation number (O.N.) of an element during oxidation?
13. What happens to the oxidation number of an element during reduction?
14. What is the oxidation state of hydrogen in hydrides?
15. What is the oxidation state of oxygen in peroxides?
16. What is the oxidation state of P_4 ?.
17. What is an electrode?
18. What is electrode potential?
18. What is standard electrode potential?
19. Name the cell obtained by coupling a zinc electrode with a copper electrode.
20. Identify the oxidant in the following reaction: $\text{H}_2\text{O}_2 + \text{O}_3 \longrightarrow \text{H}_2\text{O} + 2\text{O}_2$
21. What is the oxidation state of oxygen in OF_2 ?

QUESTIONS CARRYING TWO MARKS:

1. What is a redox reaction? Give an example.
2. Justify the reaction: $\text{H}_2\text{S} + \text{Cl}_2 \longrightarrow 2\text{HCl} + \text{S}$ is a redox reaction.
3. Define oxidation and reduction in terms of oxygen and hydrogen. Give one example for each.
4. What is oxidation number? What is the oxidation number(O.N) of Cl in KClO_3 ?
5. Define oxidation and reduction in terms of oxidation number.
6. How are the oxidizing agent and reducing agents defined in terms of oxidation number?
7. Write separate equations for the oxidation and reduction reactions occurring in the following redox reaction: $2\text{Fe} + 2\text{HCl} \longrightarrow \text{FeCl}_2 + \text{H}_2$
8. For $2\text{H}_2\text{O}_2 \longrightarrow 2\text{H}_2\text{O} + \text{O}_2$
(1) (2) (3)
 - i) What is the oxidation number of Oxygen in (2)?
 - ii) What type of Redox reaction is it?
9. Explain whether the following reaction is a redox reaction or not:
$$\text{CaCO}_3(\text{s}) \longrightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$$
- 10 Calculate the oxidation number of: (i) S in H_2SO_4 (ii) P in H_3PO_4 .
11. What is a redox couple? Identify the redox couples in the reaction:
$$\text{Zn}(\text{s}) + 2\text{Ag}^+(\text{aq}) \longrightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Ag}(\text{s})$$
- 12 What is an electrochemical series?
13. What is a spectator ion? Give an example of a reaction involving such an ion.
14. Write the formula for the following compounds represented using Stock notation:
 - (a) Nickel (II) sulphate
 - (b) Tin (IV) oxide
 - (c) Thallium (I) sulphate
 - (d) Iron (III) sulphate
15. Using Stock notation, represent the following compounds: Fe_2O_3 , CuO , MnO and MnO_2
16. Calculate the oxidation number of phosphorus in the following species:



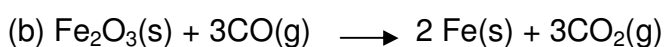
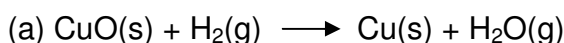
17. Balance the Redox reaction using oxidation number method :



18. Assign oxidation number to the underlined elements in each of the following

species: (a) $\text{NaH}_2\text{PO_4}$ (b) NaHSO_4 (c) $\text{H}_4\text{P2O}_7$ (d) K_2MnO_4

19. Justify that the following reactions are redox reactions:



19. Give an example of a redox combination reaction. Mention the species that undergo oxidation and reduction.

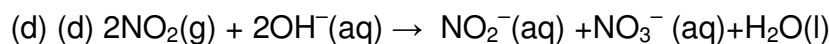
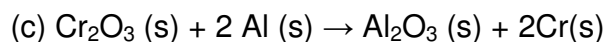
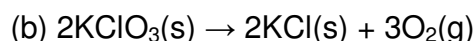
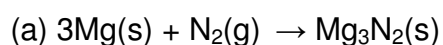
20. Give an example of a redox decomposition reaction. Mention the species that undergo oxidation and reduction.

21. Give an example of a redox displacement reaction. Mention the species that undergo oxidation and reduction.

22. Give an example of a redox disproportionation reaction. Mention the species that undergo oxidation and reduction.

23. F_2 does not undergo disproportionation. Why?

24. What type of redox reactions are the following?



25. Name the redox indicator used in the titration of

(i). KMnO_4 v/s FAS . (or $\text{H}_2\text{C}_2\text{O}_4$).

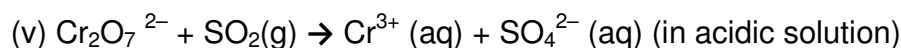
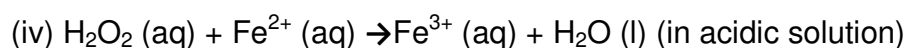
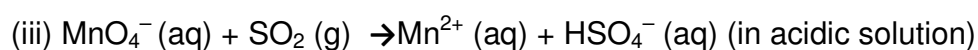
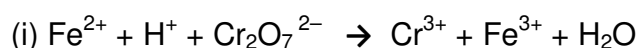
(ii) $\text{Na}_2\text{S}_2\text{O}_3$ v/s I_2 .

QUESTIONS CARRYING THREE MARKS:

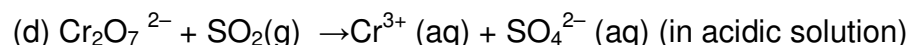
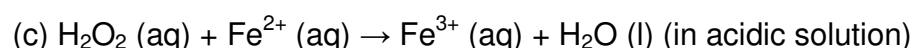
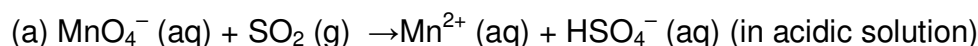
1. When blue coloured solution of copper sulphate is stirred with a zinc rod, the blue colour of the solution fades off and the zinc rod is coated with reddish copper metal. Write the chemical reaction taking place in the above observation and identify the species undergoing oxidation and reduction.

2. A solution of silver nitrate turns blue slowly on stirring with a copper rod which in turn gets coated with a white deposit of silver. Write a chemical reaction for this observation and identify the oxidizing and reducing agents in it.

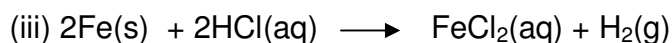
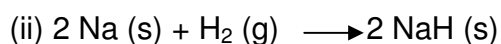
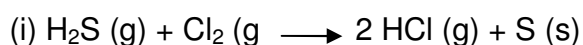
3. Balance the following equations by the oxidation number method.(3marks each)



4. Balance the following equations by half reaction method (ion-electron method). (3 marks each)



5. In the reactions given below, identify the species undergoing oxidation and reduction:



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6. Justify that the reaction: $2\text{Cu}_2\text{O}(\text{s}) + \text{Cu}_2\text{S}(\text{s}) \longrightarrow 6\text{Cu}(\text{s}) + \text{SO}_2(\text{g})$ is a redox reaction. Identify the species oxidized/reduced, which acts as an oxidant and which acts as a reductant.

CHAPTER-08 : REDOX REACTIONS

ANSWERS:

QUESTIONS CARRYING ONE MARK:

1. Loss of electron(s) by any species is called oxidation.
2. Gain of electron(s) by any species is called reduction.
3. An oxidizing agent (or an oxidant) is an acceptor of electron(s).
4. Fluorine (F₂).
5. A reducing agent(or a reductant) is a donor of electron(s).
6. Lithium (Li).
7. $2\text{Fe}^{2+} + 2\text{H}^+ + \text{H}_2\text{O}_2 \longrightarrow \underline{2\text{Fe}^{3+}} + 2\text{H}_2\text{O}$.
8. The term Oxidation number denotes the oxidation state of an element in a compound ascertained according to a set of rules formulated on the basis that electron pair in a covalent bond belongs entirely to more electronegative element.
9. Oxidation number of oxygen = -2.
Hence, oxidation number of Cr, (x) in $\text{Cr}_2\text{O}_7^{2-} = 2x + 7x(-2) = 0$, $x = +6$
10. Oxidation number of K = +1, oxygen, O = -2.
Hence, oxidation number of Mn, (x) in $\text{KMnO}_4 = (+1) + x + 4(-2) = 0$, $x = +7$
11. Zero.
12. It increases.
13. It decreases
14. In Hydrides, hydrogen has an oxidation state of -1.
15. In peroxides, oxygen has an oxidation state of -1.
16. Zero.
17. A setup consisting of a metal in contact with its salt solution is called an electrode.
18. The potential attained by a metal in contact with a solution containing its own ions is called electrode potential.

19. The potential attained by a metal in contact with its salt solution of concentration 1 mol dm^{-3} at 298 K.

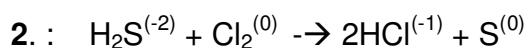
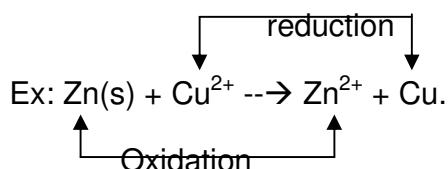
20. The oxidant is O_3 .

21. +2

QUESTIONS CARRYING TWO MARKS:

ANSWERS:

1. A chemical reaction in which both oxidation and reduction are taking place simultaneously is called a redox reaction.

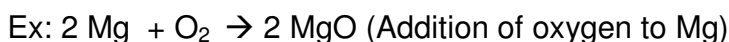


The O.N. of S increases from -2 to 0. So it is undergoing oxidation.

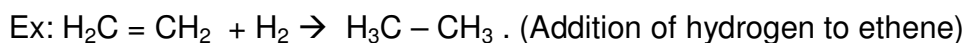
The O.N. of Cl_2 decreases from 0 to -1. So it is undergoing reduction.

Therefore it is a redox reaction.

3. **Oxidation:** Addition of oxygen or removal of hydrogen.



Reduction: Addition of hydrogen or removal of oxygen.



4. Oxidation number denotes the oxidation state of an element in a compound ascertained according to a set of rules formulated on the basis that electron pair in a covalent bond belongs entirely to more electronegative element.

Let the O.N of Cl in KClO_3 be x.

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

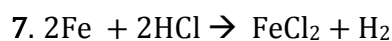
5. In terms of oxidation number,

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.

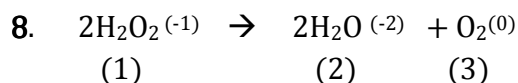
6. **Oxidising agent**: A reagent which can increase the oxidation number of an element in a given substance. These reagents are also called as **oxidants**.

Reducing agent: A reagent which lowers the oxidation number of an element in a given substance. These reagents are also called as **reductants**.



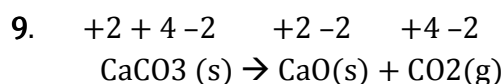
Oxidation reaction: $\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^-$

Reduction reaction: $2\text{HCl} + 2\text{e}^- \rightarrow \text{H}_2$



(i) The O.N. of oxygen in (2) is -2.

(ii) It is a disproportionation redox reaction (\because oxygen undergoes both oxidation and reduction.)



It is not a redox reaction because the oxidation number of no element changes.

10. (i) Let the O.N. of S be 'x'

$$\text{O.N. of H} = +1, \text{O} = -2 \quad \therefore \text{O.N. of S in H}_2\text{SO}_4 = 2(+1) + x + 4(-2) = +6.$$

(ii) Let the O.N. of P be 'x'.

$$\text{O.N. of H} = +1, \text{O} = -2 \quad \therefore \text{O.N. of P in H}_3\text{PO}_4 = 3(+1) + x + 4(-2) = +5.$$

11. A **redox couple** is defined as having together the oxidized and reduced forms of a substance taking part in an oxidation or reduction half reaction.

The redox couples in the reaction are, **$\text{Zn}^{2+} / \text{Zn}(\text{s})$** and **$\text{Ag}^+ / \text{Ag}$** .

12. A series of electrode potential values arranged in the increasing or decreasing order constitute an electrochemical series.

13. An ion which is present in a redox reaction, but does not take part in a reaction during electron transfer is called a spectator ion.

Ex: SO_4^{2-} ion in the reaction: $\text{Zn(s)} + \text{CuSO}_4(\text{aq}) \rightarrow \text{ZnSO}_4(\text{aq}) + \text{Cu}$.

14. (a) Ni(II)SO_4 (b) Sn(IV)O_2

(c) $\text{Tl}_2(\text{I})\text{SO}_4$ (d) $\text{Fe}_2(\text{III})(\text{SO}_4)_3$

15. $\text{Fe}_2\text{O}_3 - \text{Fe}_2(\text{III})\text{O}_3$, $\text{CuO} - \text{Cu(II)O}$

$\text{MnO} - \text{Mn(II)O}$, $\text{MnO}_2 - \text{Mn(IV)O}_2$.

16. (a) Let the O.N of P in HPO_3^{2-} be x.

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

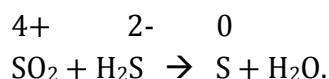
$$\therefore x = +3$$

(b) Let the O.N of P in PO_4^{3-} be x.

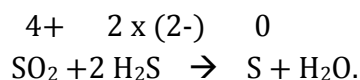
$$x + 4(-2) = -3$$

$$\therefore x = +5$$

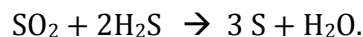
17. Step 1: Write skeletal equation with O.N of each element.



Step 2: Multiply H_2S by 2 to equalize the oxidation numbers on either side of the equation.



Step 3: Now, balance S atoms on RHS.



Step 4: Finally balance H and O atoms to get a balanced equation.



18. (a) NaH_2PO_4 : O.N. of P = $(+1) + 2(+1) + x + 4(-2)$; $x = +5$.

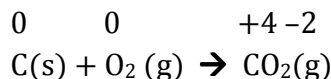
(b) NaHSO_4 : O.N of S = $(+1) + (+1) + x + 4(-2)$; $x = +6$

(c) $\text{H}_4\text{P}_2\text{O}_7$: O.N. of P = $4(+1) + 2x + 7(-2)$; $x = +5$

(d) K_2MnO_4 : O.N. of Mn = $2(+1) + X + 4(-2)$: $x = +7$

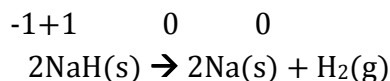
(Taking O.N. of H=+1, Na = +1, K = +1, O = -2.).

19. Example for Redox **combination** reaction:



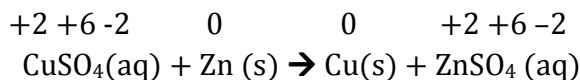
In this reaction, the O.N. of 'C' increases from 0 to +4. So it is undergoing oxidation.
the O.N. of 'O' decreases from 0 to -2. So it is undergoing reduction.

20. Example for Redox **decomposition** reaction:



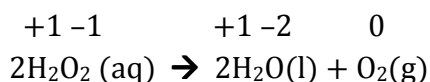
In this reaction, the O.N. of 'Na' increases from -1 to 0. So it is undergoing oxidation.
the O.N. of 'H' decreases from +1 to 0. So it is undergoing reduction.

21. Example for Redox **displacement** reaction:



In this reaction, the O.N. of 'Zn' increases from 0 to +2. So it is undergoing oxidation.
the O.N. of 'Cu' decreases from +2 to 0. So it is undergoing reduction.

22. Example for Redox **disproportionation** reaction:



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).

23. Among halogens, fluorine (F_2) is the most electronegative element; it cannot exhibit any positive oxidation state. Hence it does not show a disproportionation tendency.

24. (a) $3\text{Mg(s)} + \text{N}_2\text{(g)} \rightarrow \text{Mg}_3\text{N}_2\text{(s)}$ - Redox **combination** reaction

(b) $2\text{KClO}_3\text{(s)} \rightarrow 2\text{KCl(s)} + 3\text{O}_2\text{(g)}$ - Redox **decomposition** reaction

(c) $\text{Cr}_2\text{O}_3\text{(s)} + 2\text{Al(s)} \rightarrow \text{Al}_2\text{O}_3\text{(s)} + 2\text{Cr(s)}$ - Redox **displacement** reaction

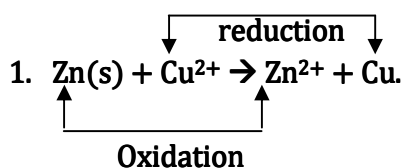
(d) $2\text{NO}_2(\text{g}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{NO}_2^-(\text{aq}) + \text{NO}_3^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$ - Redox **disproportionation** reaction.

25. (i) MnO_4^- ion itself act as a self indicator

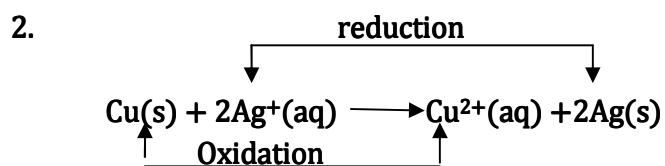
(ii) Starch.

QUESTIONS CARRYING THREE MARKS:

Answers:



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



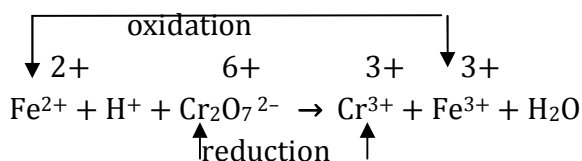
In this reaction, Cu is giving two electrons to Ag^+ and so it is a reducing agent.

Ag^+ , in turn, is accepting the electrons from Cu to undergo reduction and so it is an oxidizing agent.

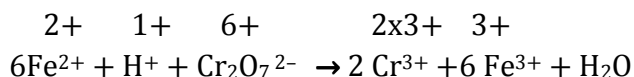
BALANCING EQUATIONS BY OXIDATION NUMBER METHOD

3. (i) $\text{Fe}^{2+} + \text{H}^+ + \text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}^{3+} + \text{Fe}^{3+} + \text{H}_2\text{O}$

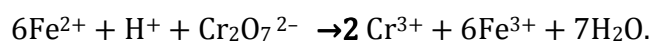
- Step 1: Write skeletal equation with O.N of each element.



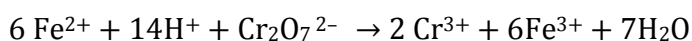
Step 2: Multiply Cr^{3+} by 2 and Fe^{2+} and Fe^{3+} by 6 to equalize the oxidation numbers on either side of the equation.



Step 3: Now, balance O atoms on RHS by adding $7\text{H}_2\text{O}$

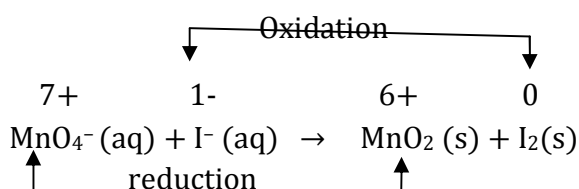


Step 4: Finally balance H atoms by adding 14H^+ on LHS to get a balanced equation as:

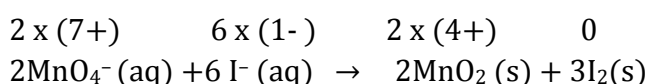


3. (ii) $\text{MnO}_4^- (\text{aq}) + \text{I}^- (\text{aq}) \rightarrow \text{MnO}_2 (\text{s}) + \text{I}_2 (\text{s})$ (in basic medium)

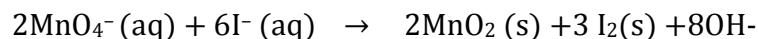
Step 1: Write skeletal equation with O.N of each element Undergoing change in oxidation number.



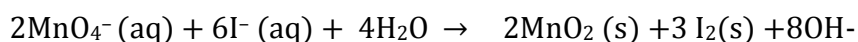
Step 2: Multiply I^- by 6 and MnO_4^- by 2 to equalize the oxidation numbers on either side of the equation.



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

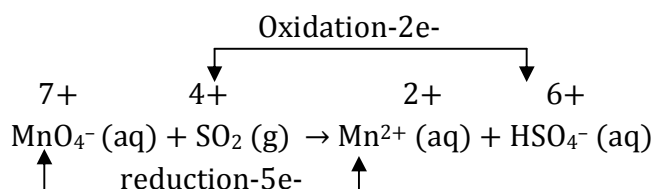


Step 4: Finally balance H and O atoms by adding $4\text{H}_2\text{O}$ on LHS to get a balanced equation as:

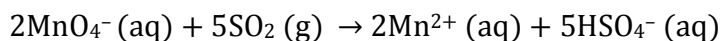


3. (iii) $\text{MnO}_4^- (\text{aq}) + \text{SO}_2 (\text{g}) \rightarrow \text{Mn}^{2+} (\text{aq}) + \text{HSO}_4^- (\text{aq})$ (in acidic solution)

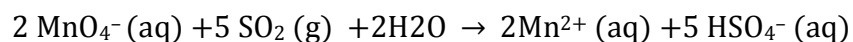
Step 1: Write skeletal equation with O.N of each element undergoing change in oxidation number.



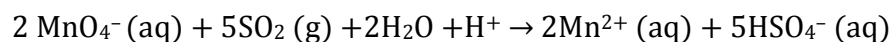
Step 2: Multiply SO_2 by 5 and MnO_4^- by 2 to balance +ve charges on both sides.



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

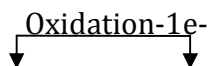


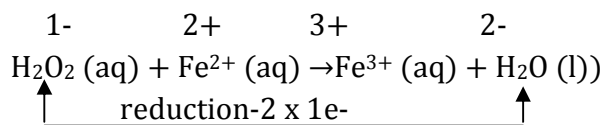
Step 4: Finally add H^+ on LHS to get a balanced equation as:



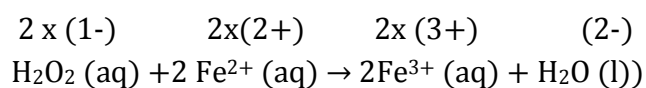
3. (iv) $\text{H}_2\text{O}_2 (\text{aq}) + \text{Fe}^{2+} (\text{aq}) \rightarrow \text{Fe}^{3+} (\text{aq}) + \text{H}_2\text{O} (\text{l})$ (in acidic solution)

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

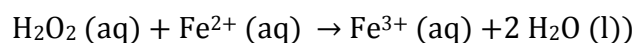




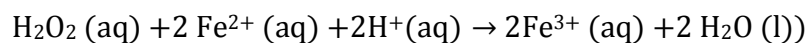
Step 2: Since the number of charges on both sides are not equal, 2Fe^{2+} on LHS and 2Fe^{3+} on RHS



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

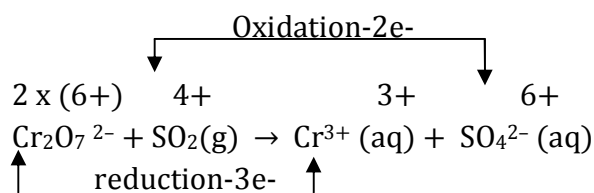


Step 4: Finally add 2H^+ on LHS to get a balanced equation as:

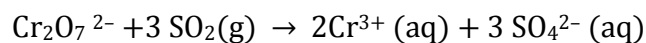


3.(v) $\text{Cr}_2\text{O}_7^{2-} + \text{SO}_2(\text{g}) \rightarrow \text{Cr}^{3+} (\text{aq}) + \text{SO}_4^{2-} (\text{aq})$ (in acidic solution)

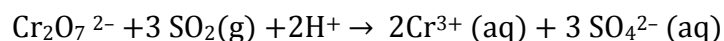
Step 1: Write skeletal equation with O.N of each element Undergoing change in oxidation number.



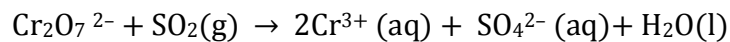
Step 2: Multiply SO_2 by 3 and Cr^{3+} by 2 on RHS .



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



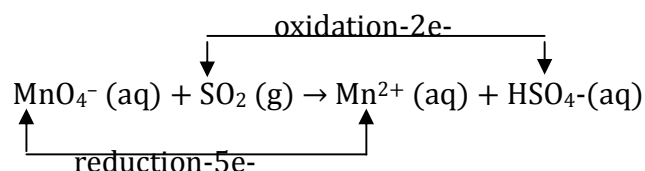
Step 4: Finally add H_2O on RHS to get a balanced equation as:



BALANCING EQUATIONS BY ION-ELECTRON METHOD

4. (a) $\text{MnO}_4^- (\text{aq}) + \text{SO}_2 (\text{g}) \rightarrow \text{Mn}^{2+} (\text{aq}) + \text{HSO}_4^- (\text{aq})$ (in acidic solution)

Step1: Assign O.N. to the atoms undergoing oxidation / reduction.



Step2: Write out oxidation and reduction separately and balance the atoms other than H and O.

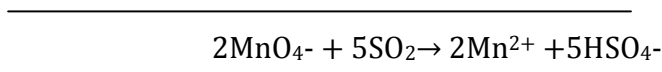
Oxidation half reaction: $\text{SO}_2 \rightarrow \text{HSO}_4^-$

Reduction half reaction: $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$

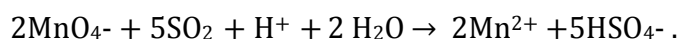
Step3: Multiply the oxidation reaction with the extent of reduction and reduction reaction by the extent of oxidation and add.

Oxidation half reaction: $[\text{SO}_2 \rightarrow \text{HSO}_4^-] \times 5$

Reduction half reaction: $[\text{MnO}_4^- \rightarrow \text{Mn}^{2+}] \times 2$

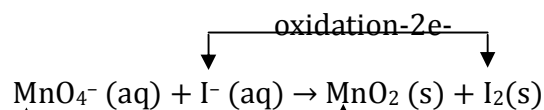


Step4: Add H^+ and $2\text{H}_2\text{O}$ on LHS to balance H and O atoms in the acid medium to get a balanced equation.



4. (b) $\text{MnO}_4^- (\text{aq}) + \text{I}^- (\text{aq}) \rightarrow \text{MnO}_2 (\text{s}) + \text{I}_2 (\text{s})$ (in basic medium)

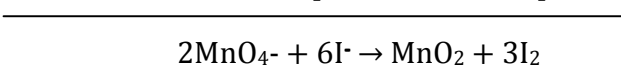
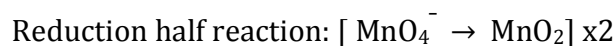
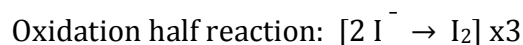
Step1: Assign O.N. to the atoms undergoing oxidation / reduction.



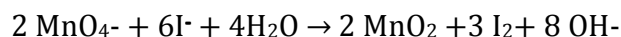
Step2: Write out oxidation and reduction separately and balance the atoms other than H and O.



Step3: Multiply the oxidation reaction with the extent of reduction and reduction reaction by the extent of oxidation and add.

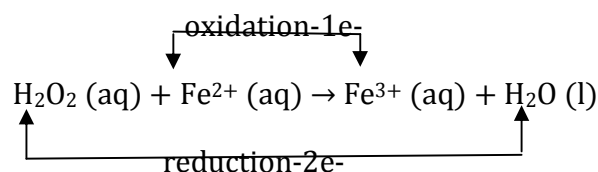


Step4: Add 4OH^- on RHS and $2\text{H}_2\text{O}$ on LHS to balance H and O atoms in the basic medium to get a balanced equation.



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

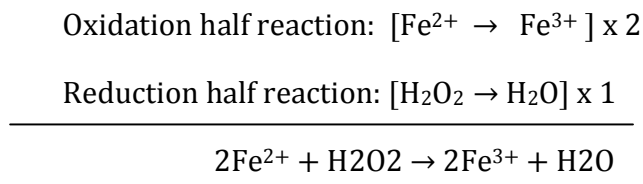
Step1: Assign O.N. to the atoms undergoing oxidation / reduction.



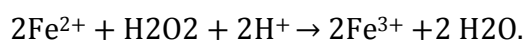
Step2: Write out oxidation and reduction separately and balance the atoms other than H and O.



Step3: Multiply the oxidation reaction with the extent of reduction and reduction reaction by the extent of oxidation and add.

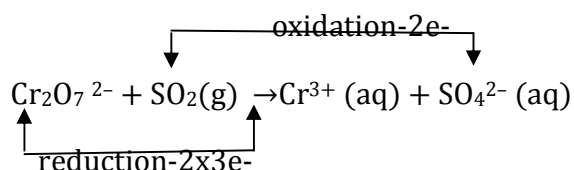


Step4: Add 2H^+ on LHS and H_2O on RHS to balance H and O atoms in the acid medium to get a balanced equation.

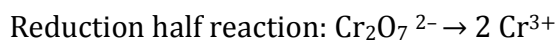


4.(d) $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + \text{SO}_2(\text{g}) \rightarrow \text{Cr}^{3+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$ (in acidic solution)

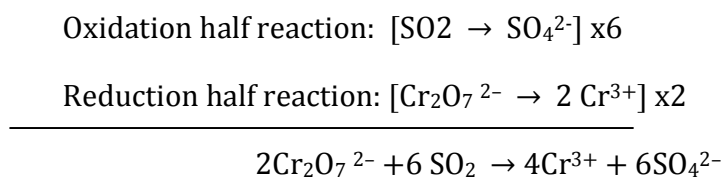
Step1: Assign O.N. to the atoms undergoing oxidation / reduction.



Step2: Write out oxidation and reduction separately and balance the atoms other than H and O.

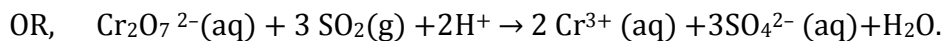
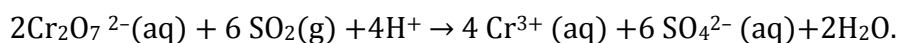


Step3: Multiply the oxidation reaction with the extent of reduction and reduction reaction by the extent of oxidation and add.



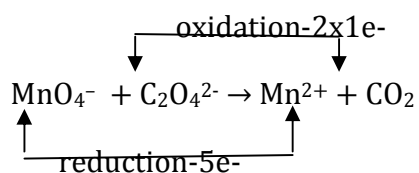
Step4: Add H^+ and $2\text{H}_2\text{O}$ on LHS to balance H and O atoms in the acid medium to get a

balanced equation.



4.(e) $\text{MnO}_4^- + \text{C}_2\text{O}_4^{2-} \rightarrow \text{Mn}^{2+} + \text{CO}_2$ in acid medium

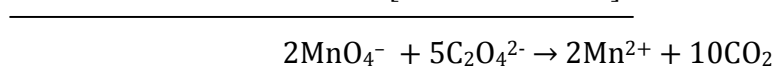
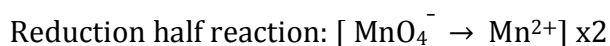
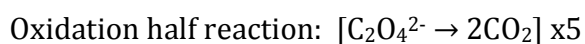
Step1: Assign O.N. to the atoms undergoing oxidation / reduction.



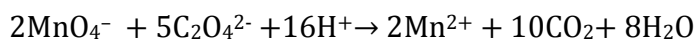
Step2: Write out oxidation and reduction separately and balance the atoms other than H and O.



Step3: Multiply the oxidation reaction with the extent of reduction and reduction reaction by the extent of oxidation and add.



Step4: 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.

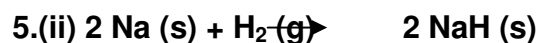


5. (i) $\text{H}_2\text{S}(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g}) + \text{S}(\text{s})$

In this reaction, the species undergoing oxidation is: H_2S

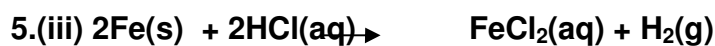
(\because the O.N. of S in H_2S increases from -2 to 0)

The species undergoing reduction is: Cl_2 (\because the O.N. of Cl decreases from 0 to -1)



In this reaction, the species undergoing oxidation is: Na (\because the O.N. of Na increases from 0 to +1).

The species undergoing reduction is: H_2 (\because the O.N. of H decreases from 0 to -1)



In this reaction, the species undergoing oxidation is: Fe (\because the O.N. of Fe increases from 0 to +2)

The species undergoing reduction is: HCl (\because the O.N. of H in HCl decreases from +1 to 0)

6. The reaction: $2\text{Cu}_2\text{O(s)} + \text{Cu}_2\text{S(s)} \rightarrow 6\text{Cu(s)} + \text{SO}_2\text{(g)}$ is a redox reaction because, in

Cu_2O , Cu is in +1 oxidation state. It is reduced to Cu in which the oxidation state is 0.

In Cu_2S , S is in -2 oxidation state, which is oxidized to +4 oxidation state in SO_2 .

The oxidizing agent(oxidant) is Cu(I) in Cu_2O

The reducing agent(reductant) is sulphur of Cu_2S .
