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: $2Fe^{2+} + 2H^{+} + H_2O_2 \longrightarrow$ .... $+2H_2O$.

8. Define oxidation number. (or oxidation state).

9. Calculate the oxidation number of Cr in $Cr_2O_7^{2-}$.

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: $H_2O_2 + O_3 \longrightarrow H_2O + 2O_2$

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: $H_2S + Cl_2 \longrightarrow 2HCl + S$ 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 if oxidation number?

7. Write separate equations for the oxidation and reduction reactions occurring in the following redox reaction: $2Fe + 2HCl \longrightarrow FeCl_2 + H_2$

8. For $\underset{(1)}{2H_2O_2} \longrightarrow \underset{(2)}{2H_2O} + \underset{(3)}{O_2}$

   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:

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

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:

$$Zn(s) + 2Ag^{+}(aq) \longrightarrow Zn^{2+}(aq) + 2Ag(s)$$

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:

(a) HPO_3^{2-} and (b)PO_4^{3-}

17. Balance the Redox reaction using oxidation number method :

$SO_2 + H_2S \longrightarrow S + H_2O$

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

species:(a) $NaH_2\underline{P}O_4$ (b) $NaH\underline{S}O_4$ (c) $H_4\underline{P}_2O_7$ (d) $K_2\underline{Mn}O_4$

19. Justify that the following reactions are redox reactions:

(a) $CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(g)$

(b) $Fe_2O_3(s) + 3CO(g) \longrightarrow 2\,Fe(s) + 3CO_2(g)$

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?

(a) $3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$

(b) $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

(c) $Cr_2O_3\,(s) + 2\,Al\,(s) \rightarrow Al_2O_3\,(s) + 2Cr(s)$

(d) (d) $2NO_2(g) + 2OH^-(aq) \rightarrow NO_2^-(aq) + NO_3^-\,(aq) + H_2O(l)$

25. Name the redox indicator used in the titration of

(i). KMnO4 v/s FAS.(or $H_2C_2O_4$).

(ii) $Na_2S_2O_3$ v/s I_2.

<u>QUESTIONS CARRYING THREE MARKS:</u>

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)

(i) $Fe^{2+} + H^{+} + Cr_2O_7^{2-} \rightarrow Cr^{3+} + Fe^{3+} + H_2O$

(ii)MnO_4^{-} (aq) + I^{-} (aq) $\rightarrow$ MnO_2 (s) + I_2(s) (in basic medium)

(iii) MnO_4^{-} (aq) + SO_2 (g) $\rightarrow Mn^{2+}$ (aq) + HSO_4^{-} (aq) (in acidic solution)

(iv) H_2O_2 (aq) + Fe^{2+} (aq) $\rightarrow Fe^{3+}$ (aq) + H_2O (l) (in acidic solution)

(v) $Cr_2O_7^{2-}$ + SO_2(g) $\rightarrow$ Cr^{3+} (aq) + SO_4^{2-} (aq) (in acidic solution)

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

(a) MnO_4^{-} (aq) + SO_2 (g) $\rightarrow Mn^{2+}$ (aq) + HSO_4^{-} (aq) (in acidic solution)

(b) MnO_4– (aq) + I– (aq) $\rightarrow$ MnO_2 (s) + I_2(s) (in basic medium)

(c) H_2O_2 (aq) + Fe^{2+} (aq) $\rightarrow$ Fe^{3+} (aq) + H_2O (l) (in acidic solution)

(d) $Cr_2O_7^{2-}$ + SO_2(g) $\rightarrow Cr^{3+}$ (aq) + SO_4^{2-} (aq) (in acidic solution)

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

(i) H_2S (g) + Cl_2 (g $\longrightarrow$ 2 HCl (g) + S (s)

(ii) 2 Na (s) + H_2 (g) $\longrightarrow$ 2 NaH (s)

(iii) 2Fe(s) + 2HCl(aq) $\longrightarrow$ $FeCl_2$(aq) + H_2(g)

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6. Justify that the reaction: $2Cu_2O(s) + Cu_2S(s) \longrightarrow 6Cu(s) + SO_2(g)$ 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 (F_2).

5. A reducing agent(or a reductant) is a donor of electron(s).

6. Lithium (Li).

7. $2Fe^{2+} + 2H^+ + H_2O_2 \longrightarrow \underline{\mathbf{2Fe^{3+}}} + 2H_2O$.

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 $Cr_2O_7^{2-}$ = 2x + 7x(-2) = 0, x = +6

10. Oxidation number of K = +1, oxygen, O = -2.

Hence, oxidation number of Mn, (x) in $KMnO_4$ = (+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 $moldm^{-3}$ 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.

reduction

Ex: $Zn(s) + Cu^{2+} \dashrightarrow Zn^{2+} + Cu$.

Oxidation

2. : $H_2S^{(-2)} + Cl_2^{(0)} \rightarrow 2HCl^{(-1)} + S^{(0)}$

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

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

Therefore it is a redox reaction.

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

Ex: $2\,Mg + O_2 \rightarrow 2\,MgO$ (Addition of oxygen to Mg)

, $2\,H_2S + O_2 \rightarrow 2\,S + 2\,H_2O$ (Removal of hydrogen from H_2S)

Reduction: Addition of hydrogen or removal of oxygen.

Ex: $H_2C = CH_2 + H_2 \rightarrow H_3C - CH_3$. (Addition of hydrogen to ethene)

$2\,HgO \rightarrow 2\,Hg + O_2$ (Removal of oxygen from HgO)

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.

O.N. of K = +1, O = -2 . $\therefore$ O.N of Cl in $KClO_3 = 1 + x + 3(-2) = +5$.

5. In terms of oxidation number,

<u>Oxidation</u>: An increase in the oxidation number of an element in a given substance.

<u>Reduction</u>: A decrease in the oxidation number of an element in a given substance.

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

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

7. $2Fe + 2HCl \rightarrow FeCl_2 + H_2$

<u>Oxidation reaction:</u> $Fe \rightarrow Fe^{2+} + 2e^-$

<u>Reduction reaction</u>: $2HCl + 2e^- \rightarrow H_2$

8. $\underset{(1)}{2H_2O_2^{(-1)}} \rightarrow \underset{(2)}{2H_2O^{(-2)}} + \underset{(3)}{O_2^{(0)}}$

(i) The O.N. of oxygen in (2) is -2.
(ii) It is a disproportionation redox reaction (∵ oxygen undergoes both oxidation and reduction.)

9. $\overset{+2\ +4\ -2}{CaCO3\ (s)} \rightarrow \overset{+2\ -2}{CaO(s)} + \overset{+4\ -2}{CO2(g)}$

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

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

O.N. of H = +1, O = -2 ∴ O.N. of S in H_2SO_4 = 2(+1) + x + 4(-2) = +6.

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

O.N. of H = +1, O = -2 ∴ O.N. of P in H_3PO_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, $\mathbf{Zn^{2+}/Zn(s)}$ and $\mathbf{Ag^+/Ag}$.

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: $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu$.

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

(c) $Tl_2(I)SO4$ (d) $Fe_2(III)(SO_4)_3$

15. Fe_2O_3 - $Fe_2(III)O_3$, CuO – $Cu(II)O$

MnO - $Mn(II)O$, MnO_2 – $Mn(IV)O_2$.

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.

$$\overset{4+}{SO_2} + H_2\overset{2-}{S} \rightarrow \overset{0}{S} + H_2O.$$

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

$$\overset{4+}{SO_2} + 2\,H_2\overset{2 \times (2-)}{S} \rightarrow \overset{0}{S} + H_2O.$$

Step 3: Now, balance S atoms on RHS.

$$SO_2 + 2H_2S \rightarrow 3\,S + H_2O.$$

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

$$\mathbf{SO_2 + 2H_2S \rightarrow 3S + 2H_2O}$$

18. (a) $NaH_2\underline{P}O_4$: O.N. of P = (+1)+2(+1)+x+4(-2); x = +5.

(b) $NaH\underline{S}O_4$: O.N of S = (+1) + (+1) + x + 4(-2); x = +6

(c) $H_4\underline{P}_2O_7$: O.N. of P = 4(+1) + 2x + 7(-2) ; x = +5

(d) $K_2\underline{Mn}O_4$: 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:

$$\overset{0}{C}(s) + \overset{0}{O_2}(g) \rightarrow \overset{+4\ -2}{CO_2}(g)$$

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:

$$2\overset{-1+1}{NaH}(s) \rightarrow 2\overset{0}{Na}(s) + \overset{0}{H_2}(g)$$

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:

$$\overset{+2\ +6\ -2}{CuSO_4}(aq) + \overset{0}{Zn}(s) \rightarrow \overset{0}{Cu}(s) + \overset{+2\ +6\ -2}{ZnSO_4}(aq)$$

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:

$$2\overset{+1\ -1}{H_2O_2}(aq) \rightarrow 2\overset{+1\ -2}{H_2O}(l) + \overset{0}{O_2}(g)$$

In this reaction, the O.N. of 'O' increases from -1 to 0 as well as decreases from -1 to -2.
So oxygen is undergoing both oxidation and reduction(disproportionation).

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) $3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$ - Redox **combination** reaction

(b) $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$ - Redox **decomposition** reaction

(c) $Cr_2O_3(s) + 2Al(s) \rightarrow Al_2O_3(s) + 2Cr(s)$ - Redox **displacement** reaction

(d) $2NO_2(g) + 2OH^-(aq) \rightarrow NO_2^-(aq) + NO_3^-(aq) + H_2O(l)$ - Redox **disproportionation** reaction.

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

(ii) Starch.

QUESTIONS CARRYING THREE MARKS:

Answers:

reduction

1. $Zn(s) + Cu^{2+} \rightarrow Zn^{2+} + Cu.$

Oxidation

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

2.

reduction

$Cu(s) + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$

Oxidation

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) $Fe^{2+} + H^+ + Cr_2O_7^{2-} \rightarrow Cr^{3+} + Fe^{3+} + H_2O$

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

oxidation

$$\overset{2+}{Fe^{2+}} + H^+ + \overset{6+}{Cr_2O_7^{2-}} \rightarrow \overset{3+}{Cr^{3+}} + \overset{3+}{Fe^{3+}} + H_2O$$

reduction

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.

$$\overset{2+}{6Fe^{2+}} + \overset{1+}{H^+} + \overset{6+}{Cr_2O_7^{2-}} \rightarrow \overset{2x3+}{2\,Cr^{3+}} + \overset{3+}{6\,Fe^{3+}} + H_2O$$

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

$$6Fe^{2+} + H^+ + Cr_2O_7^{2-} \rightarrow \mathbf{2}\,Cr^{3+} + 6Fe^{3+} + 7H_2O.$$

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

$$6\,Fe^{2+} + 14H^+ + Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+} + 6Fe^{3+} + 7H_2O$$

3. **(ii) MnO_4^- (aq) + I^- (aq) → MnO_2 (s) + I_2(s) (in basic medium)**

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

Oxidation

$$\overset{7+}{MnO_4^-}(aq) + \overset{1-}{I^-}(aq) \rightarrow \overset{6+}{MnO_2}(s) + \overset{0}{I_2}(s)$$

reduction

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

$$\overset{2 \times (7+)}{2MnO_4^-}(aq) + \overset{6 \times (1-)}{6\,I^-}(aq) \rightarrow \overset{2 \times (4+)}{2MnO_2}(s) + \overset{0}{3I_2}(s)$$

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

$$2MnO_4^- (aq) + 6I^- (aq) \rightarrow 2MnO_2 (s) + 3 I_2(s) + 8OH-$$

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

$$2MnO_4^- (aq) + 6I^- (aq) + 4H_2O \rightarrow 2MnO_2 (s) + 3 I_2(s) + 8OH-$$

3. (iii) MnO_4^- (aq) + SO_2 (g) →Mn^{2+} (aq) + HSO_4^- (aq) (in acidic solution)

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

Oxidation-2e-

$$\overset{7+}{Mn}O_4^- (aq) + \overset{4+}{S}O_2 (g) \rightarrow \overset{2+}{Mn^{2+}} (aq) + H\overset{6+}{S}O_4^- (aq)$$

reduction-5e-

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

2 x (7+) 5 x (4+) 2 x (2+) 5 x (6+)

$$2MnO_4^- (aq) + 5SO_2 (g) \rightarrow 2Mn^{2+} (aq) + 5HSO_4^- (aq)$$

Step 3: Now, add 2H2O and H+ on LHS to balance oxygen atoms

$$2 MnO_4^- (aq) + 5 SO_2 (g) + 2H2O \rightarrow 2Mn^{2+} (aq) + 5 HSO_4^- (aq)$$

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

$$2 MnO_4^- (aq) + 5SO_2 (g) + 2H_2O + H^+ \rightarrow 2Mn^{2+} (aq) + 5HSO_4^- (aq)$$

3. (iv) H_2O_2 (aq) + Fe^{2+} (aq) →Fe^{3+} (aq) + H_2O (l) (in acidic solution)

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

Oxidation-1e-

1- 2+ 3+ 2-

H_2O_2 (aq) + Fe^{2+} (aq) →Fe^{3+} (aq) + H_2O (l))

reduction-2 x 1e-

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

2Fe^{3+} on RHS

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

H_2O_2 (aq) +2 Fe^{2+} (aq) → 2Fe^{3+} (aq) + H_2O (l))

Step 3: Now, put 2H_2O to balance 'O' atoms.

H_2O_2 (aq) + Fe^{2+} (aq) → Fe^{3+} (aq) +2 H_2O (l))

Step 4: Finally add2 H^+ on LHS to get a balanced equation as:

H_2O_2 (aq) +2 Fe^{2+} (aq) +2H^+(aq) → 2Fe^{3+} (aq) +2 H_2O (l))

3.(v) $Cr_2O_7{}^{2-}$ + SO_2(g) → Cr^{3+} (aq) + $SO_4{}^{2-}$ (aq) (in acidic solution)

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

oxidation number.

Oxidation-2e-

2 x (6+) 4+ 3+ 6+

$Cr_2O_7{}^{2-}$ + SO_2(g) → Cr^{3+} (aq) + $SO_4{}^{2-}$ (aq)

reduction-3e-

Step 2: Multiply SO2 by 3 and Cr3+ by 2 on RHS .

$Cr_2O_7{}^{2-}$ +3 SO_2(g) → 2Cr^{3+} (aq) + 3 $SO_4{}^{2-}$ (aq)

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

$Cr_2O_7{}^{2-}$ +3 SO_2(g) +2H^+ → 2Cr^{3+} (aq) + 3 $SO_4{}^{2-}$ (aq)

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

$$Cr_2O_7^{2-} + SO_2(g) \rightarrow 2Cr^{3+}(aq) + SO_4^{2-}(aq) + H_2O(l)$$

<u>BALANCING EQUATIONS BY ION-ELECTRON METHOD</u>

4. **(a) MnO_4^- (aq) + SO_2 (g) $\rightarrow Mn^{2+}$ (aq) + HSO_4^- (aq) (in acidic solution)**

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

oxidation-2e-

$$MnO_4^- (aq) + SO_2 (g) \rightarrow Mn^{2+} (aq) + HSO_4^-(aq)$$

reduction-5e-

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

Oxidation half reaction: $SO2 \rightarrow HSO_4-$

Reduction half reaction: $MnO_4^- \rightarrow Mn^{2+}$

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

Oxidation half reaction: $[SO_2 \rightarrow HSO_4-]$ x5

Reduction half reaction: $[MnO_4^- \rightarrow Mn^{2+}]$ x2

$$2MnO_4- + 5SO_2 \rightarrow 2Mn^{2+} + 5HSO_4-$$

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

$$2MnO_4- + 5SO_2 + H^+ + 2\,H_2O \rightarrow 2Mn^{2+} + 5HSO_4- .$$

4. **(b) MnO_4^- (aq) + I^- (aq) $\rightarrow MnO_2$ (s) + I_2(s) (in basic medium)**

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

oxidation-2e-

$$MnO_4^- (aq) + I^- (aq) \rightarrow MnO_2 (s) + I_2(s)$$

Step2: Write out oxidation and reduction separately and balance the

atoms other than H and O.

Oxidation half reaction: $2 I^- \rightarrow I_2$

Reduction half reaction: $MnO_4^- \rightarrow MnO_2$

Step3: Multiply the oxidation reaction with the extent of reduction and reduction reaction

by the extent of oxidation and add.

Oxidation half reaction: $[2 I^- \rightarrow I_2]$ x3

Reduction half reaction: $[MnO_4^- \rightarrow MnO_2]$ x2

$$2MnO_4- + 6I^- \rightarrow MnO_2 + 3I_2$$

Step4: Add 4OH- on RHS and $2H_2O$ on LHS to balance H and O atoms in the basic medium to get a balanced equation.

$$2 MnO_4- + 6I^- + 4H_2O \rightarrow 2 MnO_2 +3 I_2+ 8 OH-$$

4. (c) H_2O_2 (aq) + Fe^{2+} (aq) → Fe^{3+} (aq) + H_2O (l) (in acidic solution)

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

oxidation-1e-

$$H_2O_2 (aq) + Fe^{2+} (aq) \rightarrow Fe^{3+} (aq) + H_2O (l)$$

reduction-2e-

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

and O.

Oxidation half reaction: $Fe^{2+} \rightarrow Fe^{3+}$

Reduction half reaction: $H_2O_2 \rightarrow H_2O$

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

Oxidation half reaction: $[Fe^{2+} \rightarrow Fe^{3+}] \times 2$

Reduction half reaction: $[H_2O_2 \rightarrow H_2O] \times 1$

$2Fe^{2+} + H2O2 \rightarrow 2Fe^{3+} + H2O$

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.

$$2Fe^{2+} + H2O2 + 2H^{+} \rightarrow 2Fe^{3+} + 2\ H2O.$$

4.**(d) $Cr_2O_7^{\ 2-}$ (aq)+ SO_2(g) $\rightarrow Cr^{3+}$ (aq) + SO_4^{2-} (aq) (in acidic solution)**

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

oxidation-2e-

$Cr_2O_7^{\ 2-} + SO_2(g) \rightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$

reduction-2x3e-

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

Oxidation half reaction: $SO_2 \rightarrow SO_4^{2-}$

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

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

Oxidation half reaction: $[SO2 \rightarrow SO_4^{2-}]$ x6

Reduction half reaction: $[Cr_2O_7^{\ 2-} \rightarrow 2\ Cr^{3+}]$ x2

$2Cr_2O_7^{\ 2-} + 6\ SO_2 \rightarrow 4Cr^{3+} + 6SO_4^{2-}$

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

$$2Cr_2O_7^{\,2-}(aq) + 6\,SO_2(g) + 4H^+ \rightarrow 4\,Cr^{3+}(aq) + 6\,SO_4^{2-}(aq) + 2H_2O.$$

OR, $$Cr_2O_7^{\,2-}(aq) + 3\,SO_2(g) + 2H^+ \rightarrow 2\,Cr^{3+}(aq) + 3SO_4^{2-}(aq) + H_2O.$$

4.(e) MnO_4^- + $C_2O_4^{2-}$ → Mn^{2+} + CO2 in acid medium

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

oxidation-2x1e-

$$MnO_4^- + C_2O_4^{2-} \rightarrow Mn^{2+} + CO_2$$

reduction-5e-

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

Oxidation half reaction: $C_2O_4^{2-} \rightarrow 2CO_2$

Reduction half reaction: $MnO_4^- \rightarrow Mn^{2+}$

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

Oxidation half reaction: $[C_2O_4^{2-} \rightarrow 2CO_2]$ x5

Reduction half reaction: $[\,MnO_4^- \rightarrow Mn^{2+}]$ x2

$$2MnO_4^- + 5C_2O_4^{2-} \rightarrow 2Mn^{2+} + 10CO_2$$

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.

$$2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$$

5. (i) H_2S (g) + Cl_2 (g) → 2 HCl (g) + S (s)

In this reaction, the species undergoing oxidation is: H_2S

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

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

5.(ii) 2 Na (s) + H_2 (g) ⟶ 2 NaH (s)

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

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

5.(iii) 2Fe(s) + 2HCl(aq) ⟶ $FeCl_2$(aq) + H_2(g)

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

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

6. The reaction: $2Cu_2O$(s) + Cu_2S(s) ⟶ 6Cu(s) + SO_2(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 Cu2S.

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