KD EDUCATION ACADEMY (9582701166

Time : 3 Hour 20 Minute | STD 11 Science Chemistry
kd 90+ ch-7 redox reaction | Total Marks : 200

* **Choose The Right Answer From The Given Options.[1 Marks Each]** **[81]**

1. The oxidation number of carbon in CH_2Cl_2 is

(A) 0 (B) 2 (C) 3 (D) 5

Ans. :

a. 0

Explanation:

Let x be the oxidation state of C in CH_2Cl_2.

Since the overall charge on the complex is 0, the sum of oxidation states of all elements in it should be equal to 0.

Therefore, $x + 2 + 2(-1) = 0$

2. he ratio of oxygen atom having −2 and −1 oxidation numbers in $S_2O_8^{2-}$ is _____.

(A) 1 (B) 2 (C) 3 (D) 4

Ans. :

c. 3

Explanation:

The ratio of oxygen atom having −2 and −1 oxidation numbers in $S_2O_8^{2-}$ is three as only one peroxy linkage is present.

So, we can see from the structure below that the number of oxygen atoms having −2 oxidation state is 6 while those having −1 oxidation state is 2.

3. Which are of the following can act as oxidising as well reducing agent?

(A) H_2 (B) I_2

(C) H_2O_2 (D) All of these

Ans. :

d. All of these

Explanation:

All of them can act as oxidising as well as reducing agent because their oxidation state can increase as well as decrease.

4. Oxidation number of Cl in $CaOCl_2$ is _____.

(A) −1 and +1 (B) +2

(C) −2 (D) None of these

Ans. :

a. −1 and +1

Explanation:

In Ca(OCl)Cl, two Cl atoms are in different oxidation state i.e., one Cl^- in -1 oxidation state and other as OCl^- in +1 oxidation state.

5. The oxidation state of C in diamond is:

(A) 0 (B) +1 (C) −1 (D) +2

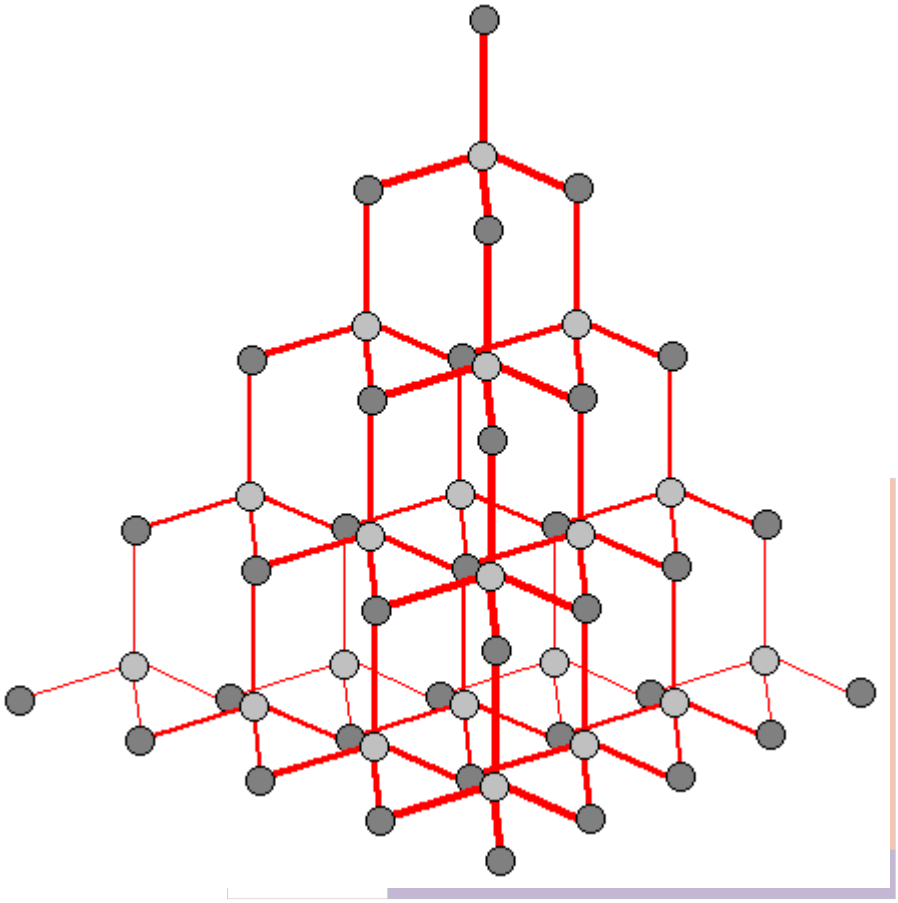
Ans. :

a. 0

Explanation:

Carbon in diamond is in elemental state, so the oxidation state of C in diamond is zero.

The structure of diamond has been shown below:



6. The oxidation number of sulphur in S_8, S_2F_2 and H_2S respectively are_____.

(A) 0, +1 and −2 (B) +2, +1 and −2

(C) 0, +1 and +2 (D) −2, +1 and −2

Ans. :

a. 0, +1 and −2

Explanation:

Oxidation number of sulphur in S_8 is 0.

Oxidation number of F is -1 so oxidation number of sulphur in S_2F_2 is +1.

Oxidation number of H is +1 so oxidation number of sulphur in H_2S is -2.

7. Identify the correct statements with reference to the given reaction:

$$P_4 + 3OH^- + 3H_2O \longrightarrow PH_3 + 3H_2PO_2^-$$

(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. Phosphorus is undergoing oxidation as well as reduction.

d. Hydrogen is undergoing neither oxidation nor reduction.

Explanation:

$$\overset{0}{P_4} + \overset{-2+1}{3OH} + \overset{+3-2}{3H_2O} \longrightarrow \overset{-3+1}{PH_3} + \overset{-2+1-2}{3H_2PO_2^-}$$

Because O.N. of P increases from 0 (P_4) to +1 ($H_2P0^-_2$) and decreases from 0 (P_4) to -3 (PH_3), therefore, P has undergone both oxidation as well as reduction. So, option (c) is correct. Option (d) is also correct because O.N. of H remains +1 in all the compounds and hence hydrogen is undergoing neither oxidation nor reduction.

8. In the balanced chemical equation:

$$IO_3 + aI^- + bH^+ \longrightarrow cH_2O + dI_2$$

a, b, c, d respectively are:

(A) 5, 6, 3, 3 (B) 5, 3, 6, 3 (C) 3, 5, 3, 6 (D) 5, 6, 5, 5

Ans. :

a. 5, 6, 3, 3

Explanation:

$$IO_3 + aI^- + bH^+ \longrightarrow cH_2O + dI_2$$

9. $H_2SO_5 + H_2O \rightarrow H_2SO_4 + H_2O_2$

Oxidation number of sulphur in H_2SO_5 in the above reaction is:

(A) 6 (B) 4 (C) 5 (D) 2

Ans. :

a. 6

Explanation:

Formation of H_2O_2 indicates that there is one peroxy linkage in H_2SO_5. The oxidation state of H = +1.

Therefore, x + 2 − 2 − 6 = 0 or, x = +6

Hence, x = 6.

Thus, oxidation number of sulphur = +6.

10. Which of the following arrangements represent increasing oxidation number of the central atom?

(A) $CrO_2^-, ClO_3^-, CrO_4^{2-}, MnO_4^-$ (B) $ClO_3^-, CrO_4^{2-}, MnO_4^-, CrO_2^-$

(C) $CrO_2^-, ClO_3^-, MnO_4^-, CrO_4^{2-}$ (D) $CrO_4^{2-}, MnO_4^-, CrO_2^-, ClO_3^-$

Ans. :

a. $CrO_2^-, ClO_3^-, CrO_4^{2-}, MnO_4^-$

Explanation:

Writing the O.N. of Cr, Cl and Mn each species in the four set of ions,

we have,

a. $\overset{+3}{Cr}\ O_2^-, \overset{+5}{Cl}\ O_3^-, \overset{+6}{Cr}\ O_4^{2-}, \overset{+7}{Mn}\ O_4^-$

b. $\overset{+5}{Cl}\ O_3^-, \overset{+6}{Cr}\ O_4^{2-}, \overset{+7}{Mn}\ O_4^-, \overset{+3}{Cr}\ O_2^-$

c. $\overset{+3}{Cr}\ O_2^-, \overset{+5}{Cl}\ O_3^-, \overset{+7}{Mn}\ O_4^-, \overset{+6}{Cr}\ O_4^{2-}$

d. $\overset{+6}{Cr}\ O_4^{2-}, \overset{+7}{Mn}\ O_4^-, \overset{+3}{Cr}\ O_2^-, \overset{+5}{Cl}\ O_3^{3-}$

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

11. The lowest possible oxidation state of nitrogen is -3 as in N^{3-}.

(A) True (B) False

(C) Ambiguous (D) None of these

Ans. :

d. None of these

Explanation:

The lowest possible oxidation state of nitrogen is -3 as in N^{3-}.

Nitrogen can form compounds in which oxidation state ranges from -3 to $+5$.

Ammonia, NH_3 and magnesium nitride, Mg_3N_2 have N in -3 oxidation state.

N has 5 valence electrons. It accepts 3 electrons to complete its octet.

12. It is found that V forms a double salt isomorphous with Mohr's salt. The oxidation number of V in this compound is_____.

(A) 3 (B) +2 (C) +4 (D) −4

Ans. :

b. +2

Explanation:

Double salt of V isomorphous with Mohr's salt is $(NH_4)\ 2V(SO_4)_2.\ 6H_2O$.

When double salt dissolves in water it dissociates to give NH_4^{+}, V^{+2} and SO4−2.

So oxidation state of vanadium is +2.

13. From the given species such as Li, K, Ca and Na, which of the following is the strongest reducing agent?

(A) Na (B) Li (C) Ca (D) K

Ans. :

b. Li

14. In $FeCr_2O_4$ the oxidation numbers of Fe and Cr are:

(A) +2 and +3 (B) 0 and +2

(C) +2 and +6 (D) +3 and +6

Ans. :

a. +2 and +3

Explanation:

$FeCr_2O_4$

$Fe: x + 6 + (4 \times -2) = 0$

$\Rightarrow x = 8 - 6$

$\Rightarrow x = +2.$

$Cr: 2 + 2x + (4 \times -2) = 0$

$\Rightarrow 2x = 8 - 2$

$\Rightarrow x = \frac{6}{2}$

$\Rightarrow x = +3.$

15. In the reaction, $2Na_2S_2O_3 + I_2 \rightarrow Na_2S_4O_6 + 2NaI$, I_2 acts as:

(A) Oxidising agent. (B) Reducing agent.

(C) Oxidising as well as reducing agent. (D) None of the above.

Ans. :

a. Oxidising agent.

16. Choose the correct explanation regarding half-reaction such as $Cr_2O_7^{2-} \longrightarrow Cr^{3+}$ from the following.

(A) It is oxidation half-reaction.

(B) Chromium being oxidized.

(C) $Cr_2O_7^{2-}$ is a good reducing agent.

(D) Chromium being reduced.

Ans. :

d. Chromium being reduced.

Explanation:

$\overset{6+}{Cr_2O_7^{2-}} \longrightarrow Cr^{3+}$

Gain of e^- (reduction)

17. Number of moles of MnO_4^- required to oxidise one mole of ferrous oxalate completely in acidic medium will be:

(A) 0.6 moles. (B) 0.4 moles.

(C) 7.5 moles. (D) 0.2 moles.

Ans. :

a. 0.6 moles.

Explanation:

$[MnO_4^- + 8H + 5C^- \longrightarrow Mn^{2+} + 3H_2O] \times 3$

$Fe^{2+} \longrightarrow Fe^{3+} + e^-$

$\begin{array}{l} COO^- \\ | \\ COO^- \end{array} \longrightarrow 2CO_2 + 2e^-$

$[Fe^{2+} + \begin{array}{l} COO^- \\ | \\ COO^- \end{array} \longrightarrow Fe^{3+} + 2CO_2 + 3e] \times 5$

$24H^+ 3MnO_4^- + 5FeC_2O_4$

$\longrightarrow 3Mn^{2+} + 5Fe^{3+} + 5CO_2 + 12H_2O$

5 moles of FeC_2O_4 is getting oxidising by 3 mole of $KMnO_4$

1 mole of FeC_2O_4 is getting oxidising by $\frac{3}{5} = 0.6$ moles

18. Which one of the following substances is a good oxidising agent?

(A) Coke.
(B) Water.
(C) Hydrogen peroxide.
(D) Sulphur dioxide.

Ans. :

c. Hydrogen peroxide.

Explanation:

A good oxidizing agent is one which can readily oxidize other chemical species and reduce itself.

Therefore, the compound which is good oxidizing agent must have initial oxidation as high, so that it can reduce its oxidation state and get reduced and oxidize others.

Coke: Reducing agent

Hydrogen peroxide (H_2O_2): because oxidation state of oxygen is -1 in hydrogen peroxide.

So, it oxidizes to 0 & acts as a reducing agent & it reduces to -2 & acts as an oxidizing agent.

But, generally, it acts as an oxidizing agent because a stable oxidation state of oxygen is -2(oxide form)

H_2O: Reducing agent.

SO_2: sulfur in the +4 oxidation state, sulfur dioxide is a reducing agent.

It is oxidized by halogens to give the sulfuryl halides, such as sulfuryl chloride.

19. Find the oxidation number of V in $Rb_4Na[HV_{10}O_{28}]$.

(A) +5
(B) +2
(C) -5
(D) None of these

Ans. :

a. +5

Explanation:

$H(1) + 1 + 1 + 10x - 2(28) = 0$

$6 + 10x - 56 = 0$

$10x = 50$

$x = +5$

20. If a reaction is carried out in acidic medium then which is used to balance the equation?

(A) H^+ ions.
(B) OH^- ions.
(C) H^- ions.
(D) O^{2-} ions.

Ans. :

a. H^+ ions.

Explanation:

If a reaction is carried out in acidic medium, H^+ ions are used to balance the equation. If it is carried out in basic medium, OH^- ions are used.

21. Oxidation number of S in $S_2O_3^{2-}$ is:

(A) -2
(B) +2
(C) +6
(D) 0

Ans. :

b. +2

Explanation:

Let Oxidation number of S in $S_2O_3^{2-}$ be x.

Thus,

$$2x + (-2 \times 3) = -2$$
$$2x - 6 = -2$$
$$2x = -2 + 6$$
$$2x = 4$$
$$x = \frac{4}{2}$$
$$x = 2$$

So the oxidation state of sulfur is +2.

22. Oxidation state of Fe in Sodium Nitroprusside is:

(A) +3 (B) +4 (C) +2 (D) +1

Ans. :

c. +2

23. $xKMnO_4 + NH_3 \longrightarrow yKNO_3 + MnO_3 + KOH + H_2O$

(A) x = 4, y = 6 (B) x = 8, y = 6

(C) x = 3, y = 8 (D) x = 8, y = 3

Ans. :

d. x = 8, y = 3

Explanation:

$$8KMnO_4 + 3NH_3 \longrightarrow 3KNO_3 + 8MnO_3 + 5KOH + 2H_2O$$

24. Consider the following chemical reaction

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

Which of the following reactions is an oxidation half-reaction?

(A) $MnO_4^-(aq) \longrightarrow MnO_2(s)$ (B) $I^-(aq) \rightarrow I_2(s)$.

(C) Both (a) and (b).

(D) None of the above.

Ans. :

b. $I^-(aq) \rightarrow I_2(s)$.

Explanation:

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

Oxidation half reaction: $I^-(aq) \longrightarrow I_2(s)$

Reduction half reaction: $MnO_4^-(aq) \longrightarrow MnO_2(s)$

25. The oxidation number of Phosphorus in $Mg_2P_2O_7$ is:

(A) +3 (B) 2 (C) +5 (D) -3

Ans. :

c. +5

Explanation:

Mg has o.n. = +2 and O has o.n. = -2

$(2 \times 2) + 2x + (7 \times -2) = 0$

$4 + 2x - 14 = 0$

$2x = 10$

$x = +5$ (o.n. of P)

26. A mole of N_2H_4 loses 10 mol of electrons to form a new compound Y. Assuming that all the nitrogen appears in the new compound, what is the oxidation state of nitrogen in Y ? (There is no change in the oxidation number of hydrogen).

(A) -1 (B) -3 (C) $+3$ (D) $+5$

Ans. :

c. $+3$

Explanation:

As given,

$N_2H_4 \longrightarrow 2y + 10e^-$

Let x be the oxidation state of N in N_2H_4.

Since, the overall charge on the complex is 0, the sum of oxidation states of all elements in it should be equal to 0.

$2x + 4 = 0$

$2y = 2x$, by replacing y, we get

$2x + 4 = 10$

or, $x = 3$

27. Oxidation state of nitrogen is not an integer in:

(A) Hydroxyl amine (NH_2OH) (B) Ammonia (NH_3)

(C) Hydrazine (N_2H_4) (D) Hydrazoic acid (N_3H)

Ans. :

d. Hydrazoic acid (N_3H)

Explanation:

In N_3H, the oxidation state of N is $-\frac{1}{3}$. Hence, it is not an integer.

The oxidation state of N in NH_2OH, NH_3 and N_2H_4 are $-1, -3$ and -2 respectively.

28. $? + O_2 \rightarrow 2K_2O$

(A) K (B) K_2 (C) 2K (D) 4K

Ans. :

d. 4K

Explanation:

$? + O_2 \rightarrow 2K_2O$

Since in the product side there are 4 atoms of potassium. so, to have a balanced equation, reactant side should also have 4K.

29. The brown ring complex compound is formulated as $[Fe(H_2O)_5(NO)]SO_4$. The oxidation state of iron in this complex is:

(A) 0 (B) +1 (C) +2 (D) +3

Ans. :

b. +1

30. The oxidation state of oxygen is maximum in_____.

(A) Bleaching powder ($CaOCl_2$)
(B) Oxygen difluoride (OF_2)
(C) Dioxygen difluoride (O_2F_2)
(D) Hydrogen peroxide (H_2O_2)

Ans. :

b. Oxygen difluoride (OF_2)

Explanation:

The oxidation states of oxygen in bleaching powder, oxygen difluoride, dioxygen difluoride and hydrogen peroxide are -2, $+2$, $+1$ and -1 respectively.

31. The oxidation number of cobalt in $K[Co(CO)_4]$ is:

(A) -1 (B) -3 (C) $+1$ (D) $+3$

Ans. :

a. -1

Explanation:

The oxidation number of colbalt in $K[Co(CO)_4]$ is:

$+1 + x + 0 \times 4 = 0$

$\therefore x = -1.$

32. Which of the following processes does not involve oxidation of iron?

(A) Rusting of iron sheets.
(B) Decolourisation of blue $CuSO_4$ solution by Fe.
(C) Formation of $Fe(CO)_5$ from Fe.
(D) Liberation of H_2 from steam by iron at high temperature.

Ans. :

c. Formation of $Fe(CO)_5$ from Fe.

Explanation:

$\because$ There is no change in oxidation state.

$$\underset{0}{Fe} + \underset{0}{5CO} \longrightarrow \underset{0}{Fe(CO)_5}$$

33. The value of n in the molecular formula $Be_nAl_2Si_6O_{18}$ is:

(A) 1 (B) 2 (C) 3 (D) 4

Ans. :

c. 3

34. In the reaction between copper nitrate solution and zinc, copper ions are reduced by gaining electrons from:

(A) Copper. (B) Nitrogen. (C) Zinc. (D) Oxygen.

Ans. :

c. Zinc.

35. When ammonium nitrate is gently heated, an oxide of nitrogen is formed. What is the oxidation state of nitrogen in this oxide?

(A) +4 (B) +2 (C) +3 (D) +1

Ans. :

d. +1

Explanation:

When ammonium nitrate is gently heated, dinitrogen monoxide N_2O is obtained.

Let x be the oxidation state of N in N_2O.

The oxidation state of oxygen is −2.

The sum of the oxidation states for neutral molecule is 0.

Hence, 2x+(−2) = 0 or, x = +1.

Therefore, the oxidation state of N is +1.

36. An element if present in the free or the uncombined state, its each atom bears an oxidation number:

(A) More than 1 (B) Less than 1

(C) More than 2 (D) Zero.

Ans. :

d. Zero.

37. The oxidation number of P in $Na_4P_2O_7$ is:

(A) +3 (B) +2 (C) +5 (D) −3

Ans. :

c. +5

38. The average oxidation number of iodine in I_3^- ion is:

(A) -1 (B) $\frac{-1}{3}$ (C) $+1$ (D) $\frac{+1}{3}$

Ans. :

b. $\frac{-1}{3}$

Explanation:

I_3^-

$\because 3x = 1$

$\Rightarrow x = \frac{-1}{3}$

39. Oxidation number of sulphur in marshall's acid ($H_2S_2O_8$) is:

(A) +5 (B) +8 (C) +6 (D) +7

Ans. :

c. +6

Explanation:

Let x be the oxidation number of sulphur in marshall's acid ($H_2S_2O_8$). It contains a peroxide linkage. Thus, the oxidation number of two O atoms is −1 and that of the remaining O atoms is −2. The oxidation number of H is +1.

For the neutral molecule, the sum of the oxidation numbers is zero.

Therefore, 2(+1) + 2x + 2(−1) + 6(−2) = 0 or, x = +6

40. $E^{\ominus}$ values of some redox couples are given below. On the basis of these values choose the correct option:

$$E^{\ominus}\text{values}: Br_2/\,Br^- = +1.90;\ Ag^+/\,Ag(S) = +0.80$$
$$Cu^{2+}/\,Cu(s) = +0.34; I_2(s)/\,I^- = 0.54$$

(A) Cu will reduce Br^- (B) Cu will reduce Ag

(C) Cu will reduce I^-

(D) Cu will reduce Br_2

Ans. :

d. Cu will reduce Br_2

Explanation:

Copper will reduce Br_2, if the E° of the redox reaction, 2Cu + Br_2 $CuBr_2$ is +ve.

Now $Cu \longrightarrow Cu^{2+} + 2e^-; E^0 = -0.34V$

$Br_2 + 2e^- \longrightarrow 2Br^-; E^0 = +1.90V$

$Cu + Br_2 \longrightarrow CuBr_2; E^0 = +1.56V$

Since E° of this reaction is +ve, therefore, Cu can reduce Br2 and hence option (d) is correct.

41. Oxidation number of C in HCCOH is_____.

(A) +2 (B) +4 (C) +3 (D) 0

Ans. :

a. +2

42. One mole of N_2H_4 loses 10 moles of electrons to form a new compound A. Assuming that all the nitrogen appears in the new compound, what is the oxidation state of nitrogen in A?

[There is no change in the oxidation state of hydrogen]

(A) +1 (B) −3 (C) +3 (D) +5

Ans. :

c. +3

Explanation:

The oxidation state of N in hydrazine is -2.

1 mole of hydrazine contains 2 moles of N and loses 10 moles of electrons.

Hence, 1 N atom will lose 5 electrons.

Hence, its oxidation number will increase by 5.

Hence, the oxidation number of N in compound A will be −2 + 5 = +3.

43. The oxidation number of Mn in potassium permanganate is:

(A) +6 (B) +7 (C) +5 (D) +8

Ans. :

b. +7

Explanation:

Let x be the oxidation number of manganese in potassium permanganate. ($KMnO_4$)

The oxidation numbers of potassium and oxygen are +1 and −2 respectively.

The sum of the oxidation numbers in a neutral molecule is zero.

44. On reduction of $KMnO_4$ by oxalic acid in acidic medium, the oxidation number of Mn chnages. What is the magnitude of this change?

(A) 7 to 2 (B) 6 to 2 (C) 5 to 2 (D) 7 to 4

Ans. :

a. 7 to 2

Explanation:

In acidic medium reduction of $KMnO_4$ takes place as follows:

So oxidation state of Mn changes from +7 to +2.

45. 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:

$E^{\ominus}$ values: $Fe^{3+}/\ Fe^{2+} = +0.77; I_2(S)/\ I^- = +0.54;$

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

(A) Fe^{3+} (B) $I_2(S)$ (C) Cu^{2+} (D) Ag^+

Ans. :

d. Ag^+

Explanation:

Since $Ag^+/\ Ag$ has highest positive value of $E^{\ominus}$, therefore, Ag^+ is the strongest oxidizing agent with highest tendency to get reduced.

46. What is the oxidation number of chlorine in ClO_3^-?

(A) +5 (B) +3 (C) +4 (D) +2

Ans. :

a. +5

Explanation:

Let x be the oxidation number of chlorine in ClO_3^-

The oxidation number of oxygen is -2.

The sum of the oxidation numbers of chlorine and oxygen is -1, which is equal to the charge on ion.

47. Solution of potassium chloride or ammonium nitrate in salt-bridge usually solidified by boiling with:

(A) Agar-agar. (B) Starch. (C) Cellulose. (D) Glycogen.

Ans. :

a. Agar-agar.

Explanation:

A solution of potassium chloride or ammonium nitrate is solidified by boiling with agar-agar and later cooling to a jelly like substance.

48. In the reaction, $2KClO_3 \rightarrow 2KCl + 3O_2$, the elements which have been oxidised and reduced respectively are:

(A) Chlorine and oxygen. (B) Oxygen and chlorine.

(C) Potassium and oxygen. (D) Oxygen and potassium.

Ans. :

b. Oxygen and chlorine.

49. What is the oxidation number of Br in the compound $RbBrO_4$?

(A) -1 (B) +7 (C) +1 (D) +4

Ans. :

b. +7

Explanation:

$RbBrO_4$

Total charge = 0.

Rb is a group I element. So, its oxidation no. = +1.

Oxidation no. of oxygen = −2.

Let, Oxidation no. of Br be x.

So, 1 + x + 4(−2) = 0

1 + x − 8 = 0

x − 7 = 0

x =7

50. When tin(IV) chloride is treated with excess of conc. hydrochloric acid, the complex ion$(SnCl_6)^{2-}$ is formed. The oxidation state of tin in this complex ion is?

(A) +4 (B) zero (C) −2 (D) −4

Ans. :

a. +4

Explanation:

Let oxidation state of Sn is x and we know oxidation of Cl is −1, so x + 6(−1) = −2, x = +4.

51. Which of the following compounds we use in our laboratory as a standard solution (titrant) ?

(A) $KMnO_4$ (B) $K_2Cr_2O_7$

(C) $Na_2S_2O_3$ (D) All of these

Ans. :

d. All of these

Explanation:

A reagent, called the titrant or titrator is prepared as a standard solution. A known concentration and volume of titrant reacts with a solution of analyte or titrand to determine concentration.

$KMnO_4$, $K_2Cr_2O_7$, $Na_2S_2O_3$ etc. are compounds we use in our laboratory as a standard solution.

52. In the chemical reaction,

$K_2Cr_2O_7 + xH_2SO_4 + ySO_2 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + zH_2O$

the value of x + y + z

(A) 6 (B) 5 (C) 7 (D) 3

Ans. :

b. 5

Explanation:

The balanced redox reaction is:

$K_2Cr_2O_7 + H_2SO_4 + 3SO_2 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + zH_2O$

x = 1, y = 3, z = 1

53.

The difference in the oxidation numbers of the two types of sulphur atoms in $Na_2S_4O_6$ is:

(A) 5 (B) 4 (C) 3 (D) 6

Ans. :

a. 5

Explanation:

In $Na_2S_4O_6$, the oxidation number of end sulphur atoms is +5 each and the oxidation number of middle sulphur atoms is 0 each.

The difference in the oxidation numbers of the two types of sulphur atoms is $5 - 0 = 0$.

54. What is the oxidation number of Si in the compound $CaSiO_3$?

(A) −4 (B) +2 (C) −2 (D) +4

Ans. :

d. +4

Explanation:

$CaSiO_3$

Total charge present = 0.

Oxidation no. of Oxygen is −2.

Oxidation no. od Calcium is +2.

Let, oxidation no. of Slilicon be X.

So, [+2] + x + 3[−2] = 0

2 + x − 6 = 0

x = 4

55. Standard reduction frotential of X, Y, Z are -1.2v, +0.5v, -3.0v respectively, the reducing power of the metals will be:

(A) Y > Z > X (B) Y > X > Z

(C) Z > X > Y (D) X > Y > Z

Ans. :

a. Y > Z > X

Explanation:

'Z' is best because it has lowest standard reduction potential whereas 'Y' is weakest due to highest standard reduction potential.

56. When P reacts with caustic soda, the products are PH_3 and NaH_2PO_2. The reaction is an example of.

(A) Oxidation. (B) Reduction.

(C) Both oxidation and reduction. (D) Neutralisation.

Ans. :

c. Both oxidation and reduction.

Explanation:

$P_4 + 3NaOH + 3H_2O \rightarrow PH_3 + 3NaH_2PO_2$

In reactant P is present in (0) oxidation state and in PH_3, it is present in (-3) oxidation state and in NaH_2PO_2 it is present in (+1) oxidation state.

57. The oxidation state of the most electronegative element in the products of the reaction between BaO_2 and H_2SO_4 are:

(A) 0 and −1 (B) −1 and −2

(C) −2 and 0 (D) −2 and +1

Ans. :

b. −1 and −2

Explanation:

$BaO_2 + H_2SO_4 \rightarrow BaSO_4 + H_2O_2$

The most electronegative element in the product is oxygen.

The oxidation state of oxygen in $BaSO_4$ is -2 and in H_2O_2 is -1.

58. Tailing of mercury is _____ redox change.

(A) Intramolecular. (B) Intermolecular.

(C) Disproportion. (D) None.

Ans. :

b. Intermolecular.

Explanation:

When ozone is passed through mercury, mercurous oxide (Hg2O) is formed.

Due to this, mercury loses its meniscus and starts sticking to the glass.

This phenomenon is known as Tailing of mercury.

$2Hg + O_3 \rightarrow Hg_2O + O_2$.

In this reaction, the oxidation number of mercury changes from 0 to +1. Thus, it is oxidized.

The oxidation number of oxygen changes from 0 to -2. Thus, it is reduced.

59. In the given reaction,

$CH_2 = CH_2(g) + H_2(g) \rightarrow CH_3—CH_3(g)$

ethene undergoes:

(A) Reduction process. (B) Oxidation process.

(C) Addition process. (D) All of these.

Ans. :

a. Reduction process.

Explanation:

$CH_2=CH_2 + H—H \rightarrow H_3C—CH_3$

(Addition of hydrogen)

Because of the addition of hydrogen, there occurs reduction of ethylene.

60. The oxidation state of Cr in $K_2Cr_2O_7$ is:

(A) +4 (B) +3 (C) +6 (D) +5

Ans. :

c. +6

Explanation:

$K_2Cr_2O_7$

Let the oxidation state of Cr is x.

$2(+1) + 2x + 7(-2) = 0$

$+2 + 2x - 14 = 0$

$2x - 12 = 0$

$2x = 12$

$x = +6$

61. In $MgCl_2$, the oxidation number of chlorine is:

(A) +1 (B) +2 (C) −1 (D) 0

Ans. :

c. −1

Explanation:

In $MgCl_2$, oxidation number of Cl is:

$\Rightarrow 2 + 2x = 0$

$\Rightarrow x = -1.$

62. The sum of oxidation number of all the atoms in a neutral molecule must be zero.

(A) True. (B) False.

(C) Ambiguous. (D) None of these.

Ans. :

a. True.

Explanation:

The sum of oxidation number of all the atoms in a neutral molecule must be zero.

For example, neutral molecules such as O_2, P_4, O_3, S_8 and $KMnO_4$ have the sum of oxidation number of all the atoms equal to zero.

For an ion, the sum of oxidation number of all the atoms is equal to the charge on the ion.

For example, in cyanide ion (CN^-), the sum of oxidation number of all the atoms is equal to −1.

In ammonium ion $(NH_4)+$, the sum of oxidation number of all the atoms is equal to +1.

63. The oxidation number of chromium in CrO_5 is:

(A) +6 (B) +5 (C) +10 (D) 0

Ans. :

a. +6

Explanation:

The oxidation number of chromium in chromium pentaoxide is 6.

64. Oxidation state of nitrogen in NH_2OH is:

(A) −3 (B) −1 (C) +2 (D) 3

Ans. :

b. −1

Explanation:

Let x be the oxidation state of N in NH_2OH.

Since the overall charge on the complex is 0, the sum of oxidation states of all elements in it should be equal to 0.

Therefore, $x + 2 - 1 = 0$ or, $x = -1$.

65. Oxygen has an oxidation state of +2 in.

(A) H_2O_2 (B) OF_2 (C) SO_2 (D) H_2O

Ans. :

b. OF_2

Explanation:

Oxidation state of oxygen is always -2 except in peroxides,superoxides and when it reacts with fluorine.

In H_2O_2, oxidation state of H is +1, so oxidation state of oxygen is -1.

In OF_2, oxidation state of F is -1, so oxidation state of oxygen is +2.

In SO_2 and H_2O, oxidation state of oxygen is -2.

66. The oxidation numbers of sulphur in S_8, S_2F_2 and F_2S respectively, are:

(A) 0, +1 and −2 (B) +2, +1 and −2

(C) 0, +1 and +2 (D) −2, +1 and −2

Ans. :

c. 0, +1 and +2

67. In which of the following compounds, is the oxidation number of sulphur is the least?

(A) SO_2 (B) SO_3 (C) $Na_2S_4O_8$ (D) H_2SO_4

Ans. :

c. $Na_2S_4O_8$

Explanation:

The oxidation number of sulphur in SO_2,SO_3,Na_2SO_4 and H_2SO_4 are +4,+6,+2.5 and +6 respectively.

68. The oxidation numbers of the sulphur atoms in peroxy monosulphuric acid (H_2SO_5) and peroxydisulphuric and ($H_2S_2O_8$) are respectively.

(A) +8 and +7 (B) +3 and +3

(C) +6 and +6 (D) +4 and +6

Ans. :

c. +6 and +6

Explanation:

By looking the structure of H_2SO_5, we can observe that there are two oxygen atoms which are linked by peroxide linkage so their oxidation numbers are -1.Rest oxygen atoms attached normally so their oxidation state is -2.

The oxidation number of hydrogen is +1.So oxidation number of sulphur is

$2(+1) + x + 2(-1) + 3(-2) = 0$, $x = +6$

69. The oxidation state of the underlined element in the given compound is:

BaCl2

(A) +2 (B) -2

(C) 0 (D) None of these

Ans. :

a. +2

Explanation:

$BaCl_2$

$\Rightarrow x + (-2) = 0$

$\Rightarrow x = 2$

As chlorine needs only one electron to get octet.

70. What is the oxidation state of central atom in $Ca[PtCl_4]$?

(A) 1 (B) 2 (C) 3 (D) 4

Ans. :

b. 2

Explanation:

$Ca[PtCl_4] \Leftrightarrow Ca^{+2}[PtCl4]^{2-}$

Take $[PtCl4]^{2-}$. Central atom = Pt.

Let x be the oxidation no. of Pt,

Cl oxidation no. is -1.

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

$x - 4 = -2$

$x = 2$

71. Which among the following shows maximum oxidation state?

(A) V (B) Fe (C) Mn (D) Cr

Ans. :

c. Mn

Explanation:

Metal	Maximum Oxidation state
V	+3
Cr	+6
Fe	+3
Mn	+7

72. What is the oxidation number of O in a diatomic molecule (O_2)?

(A) +2 (B) ±2 (C) +8 (D) 0

Ans. :

d. 0

Explanation:

The oxidation state of any element in its native state is zero.

73. In which of the following, the highest oxidation state is not possible?

(A) $[XeO_6]^{4-}$ (B) XeF_8 (C) OsO_4 (D) RuO_4

Ans. :

b. XeF_8

Explanation:

Xe shows +8 oxidation state in XeF_8 but it does not exist because of steric hindrance of 8F atoms.

74. What is the oxidation state of Mn in the compound K_2MnO_4?

(A) 3 (B) 4 (C) 5 (D) 6

Ans. :

d. 6

75. Plumbous ion is represented as:

(A) Pb^{+2} (B) Pb^{+4} (C) Pb^{+3} (D) Pb^{+1}

Ans. :

a. Pb^{+2}

Explanation:

Pb^{2+} = Plumbous ion

Pb^{4+} = Plumbic ion

76. Oxidation number of C in HNC is_____.

(A) +2 (B) −3 (C) +3 (D) Zero

Ans. :

a. +2

Explanation:

Oxidation number of hydrogen is +1. As nitrogen is more electronegative than carbon so its oxidation number is -3. Net charge on compound is zero.

Now we can find oxidation number of carbon:

Let oxidation number of carbon is x,

1 + (−3) + x = 0

x = +2

77. In which of the following groups of iodine compounds shows increasing order of oxidation states:

(A) HIO_4, ICl, I_2, HI (B) HI, I_2, IC, HIO_4

(C) I_2, HI, HIO_4 HI (D) ICl HIO_4,HI, I_2

Ans. :

a. HI, I_2, IC, HIO_4

Explanation:

HI(-1), I_2(O), ICl(+1), HIO_4(+7)

78. Which of the following is not an example of redox reaction?

(A) $CuO + H_2 \longrightarrow Cu + H_2O$

(B) $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$

(C) $2K + F_2 \longrightarrow 2KF$

(D) $BaCl_2 + H_2SO_4 \longrightarrow BaSO_4 + 2HCl$

Ans. :

d. $BaCl_2 + H_2SO_4 \longrightarrow BaSO_4 + 2HCl$

Explanation:

$BaCl_2 + H_2SO_4 \longrightarrow BaSO_4 + 2HCl$ is not a redox reaction. It is an example of double displacement reactions.

79. Identify the correct statement (s) in relation to the following reaction:

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

(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. Hydrogen ion is acting as an oxidant.
d. Zinc is acting as a reductant.

Explanation:

$$\overset{0}{Zn} + 2\,\overset{+1-1}{HCl} \longrightarrow \overset{+2-1}{ZnCl_2} + \overset{0}{H_2}$$

a. The O.N. of Zn increases from 0 to +2 (in $ZnCl_2$) and therefore, Zn acts as a reductant and not as an oxidant. Hence, option (a) is not correct.
b. The O.N. of Cl does not change and therefore, it neither acts as a reductant nor an oxidant. Hence, option (b) is not correct.
c. The O.N. of H decreases from +1 in H^+ to 0 in H_2. Therefore, H^+ acts an oxidant. This option is correct.
d. Zinc acts as reductant because its O.N. changes from 0 to +2. This option is correct.

80. In $Ni(CO)_4$, the oxidation state of Ni is?

(A) 4 (B) Zero (C) 2 (D) 8

Ans. :

a. Zero

Explanation:

In nickel tetracarbonyl, the oxidation state for nickel is assigned as zero. The formula conforms to the 18-electron rule. The molecule is tetrahedral, with four carbonyl (carbon monoxide) ligands attached to nickel.

81. What is the oxidation number of lithium in LiCl?

(A) +3 (B) −1 (C) +1 (D) 0

Ans. :

c. +1

Explanation:

Oxidation number of Li in LiCl: x − 1 = 0

⇒ x = +1

* Answer The Following Questions In One Sentence.[1 Marks Each] [11]

82. Justify that the following reactions are redox reactions:

$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$

Ans. : $CuO_{(s)} + H_{2(g)} \rightarrow Cu_{(s)} + H_2O_{(g)}$

Let us write the oxidation number of each element involved in the given reaction as:

$$\overset{+2\ -2}{Cu\ O_{(s)}} + \quad \overset{0}{H_{2(g)}} \quad \rightarrow \quad \overset{0}{Cu_{(s)}} + \quad \overset{+1\ \ -2}{H_2\ \ O_{(g)}}$$

Here, the oxidation number of Cu decreases from +2 in CuO to 0 in Cu i.e., CuO is reduced to Cu. Also, the oxidation number of H increases from 0 in H_2 to +1 in H_2O i.e., H_2 is oxidized to H_2O. Hence, this reaction is a redox reaction.

83. What happens when Cu^{2+} is added KI solution? Indicator used in this titration?

Ans. : $2Cu^{2+} + 4I^{-}(aq) \longrightarrow Cu_2I_2(s) + I_2(aq)$

Starch is used as indicator which gives blue colour with I_2.

84. What is the relationship between direction of current and flow of electrons by convention?

Ans. : The current flows from cathode to anode, whereas electrons flow from anode to cathode.

85. What is oxidation state of Cr in $[Cr(H_2O)_6]Cl_3$

Ans. : Let oxidation state of Cr be 'x', H is +1, O is -2,

Cl = -1

x + 12 - 12 - 3 = 0

x = 3

86. What are spectator ions? Give one example.

Ans. : Spectator ions are ions that stay unaffected during a chemical reaction. They appear both as reactant and as product in an ionic equation, e.g. in the following ionic equation, the sodium and nitrate ions are spectator ions.

$$Ag^{+}(aq) + NO_3^{-}(aq) + Na^{+}(aq) + Cl^{-}(aq) \longrightarrow AgCl(s) + Na^{+}(aq) + NO_3^{-}(aq)$$

87. What is the relationship between standard oxidation potential and standard reduction potential?

Ans. : Both are equal in magnitude but opposite in sign.

88. Out of Zn and Cu vessel which one will be more suitable to store 1M HCl?

$E^{\circ}_{\frac{Zn^{2+}}{Zn}} = -0.76V$

$E^{\circ}_{\frac{Cu^{2+}}{Cu}} = +0.34V$

Ans. : 'Cu' vessel is more suitable because Cu is less reactive than hydrogen due to higher value of reduction potential where 'Zn' is more reactive than hydrogen, will displace H_2 from IM HCl.

89. How to find strength of $KMnO_4$ by titrating it with Mohr's salt in acidic medium?

Ans. : $5M_1V_1 = M_2V_2$ is used because in $KMnO_4$, Mn^{7+} changes ($KMnO_4$) (Mohr's salt) to Mn^{2+} by gaining 5 electrons, therefore we have $5M_1V_1$ but in Mohr's salt ($FeSO_4(NH_4).6H_2O$, Fe^{2+} loses one electrons to form Fe^{2+} therefore M_2V_2 is used.

90. Justify that the following reactions are redox reactions:

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

Ans. : $Fe_2O_{3(s)} + 3CO_{(g)} \rightarrow 2Fe_{(s)} + 3CO_{2(g)}$

Let us write the oxidation number of each element in the given reaction as:

$$\overset{+3}{Fe_2}\overset{-2}{O_{3(s)}} + 3\overset{+2-2}{CO_{(g)}} \rightarrow \overset{0}{2Fe_{(s)}} + 3\overset{+4-2}{CO_{2(g)}}$$

Here, the oxidation number of Fe decreases from +3 in Fe_2O_3 to 0 in Fe i.e., Fe_2O_3 is reduced to Fe. On the other hand, the oxidation number of C increases from +2 in CO to +4 in CO_2 i.e., CO is oxidized to CO_2. Hence, the given reaction is a redox reaction.

91. $Br_2 + 2Cl^- \rightarrow Cl_2 + 2Br^-$, will this reaction take place or not?

$$E^0_{\frac{Br_2}{Br^-}} = +1.09V$$

$$E^0_{\frac{Cl_2}{Cl^-}} = +1.36V$$

Ans. : $E^0_{cell} = E^0_{\frac{Br_2}{Br^-}} - E^0_{\frac{Cl_2}{Cl^-}}$

$= 1.09V - 1.36$

$= -0.27V$

Since E^o cell is -ve, reaction will not take place.

92. Refer to the periodic table given in your book and now answer the following questions:
Select the possible non metals that can show disproportionation reaction.

Ans. : In disproportionation reactions, one of the reacting substances always contains an element that can exist in at least three oxidation states.

P, Cl, and S can show disproportionation reactions as these elements can exist in three or more oxidation states.

*** Given Section consists of questions of 2 marks each.** **[22]**

93. Assign oxidation number to the underlined elements in the following species:
$H_4\underline{P_2}O_7$

Ans. : $\overset{+1}{H_4}\quad \overset{x}{P_2}\quad \overset{-2}{O_7}$

Then, we have

4(+1) + 2(x) + 7(-2) = 0

⇒ 4 + 2x - 14 = 0

⇒ 2x = +10

⇒ x = +5

Hence, the oxidation number of P is +5.

94. The compound Y $Ba_2Cu_3O_7$, which shows superconductivity, has copper in x oxidation state. Assume that the rare earth element yttrium is in its usual +3 oxidation state. Predict the value of x.

Ans. : 1x (+3) + 2x (+ 2) + 3x + 7x (-2) = 0

3 + 4 + 3x - 14 = 0

3x = 7; $x = \frac{7}{3}$

95. Permanganate ion reacts with bromide ion in basic medium to give manganese dioxide and bromate ion. Write the balanced chemical equation for the reaction.

Ans. : Step 1:

$$\underset{+7}{MnO_4^-}(aq) + \underset{-1}{Br}(aq) \longrightarrow \underset{+4}{MnO_2}(aq) + \underset{+5}{BrO_3^-}(aq)$$

Step 2: MnO_4^- is oxidant because its oxidation state is decreasing. Br^- is reductant because its oxidation state is increasing.

Step 3: Oxidation state of Mn is decreasing by 3. Oxidation state of Br is increasing by 6. To equalize increase and decrease, multiply MnO_4^- by 2 and Br^- by 1 we get.

$$2MnO_4^-(aq) + Br^-(aq) \longrightarrow MnO_2(s) + Br_2^-(aq)$$

Step 4: Now for balncaing oxygen, we add 1 molecule of H_2O on RHS.

$$2MnO_4^-(aq) + Br^-(aq)$$
$$\longrightarrow MnO_2(s) + BrO_3^-(aq) + H_2O(l)$$

Step 5: As the reaction is taking place in basic medium to balance hydrogen, add $2H_2O$ molecules on LHS and $2OH^-$ on RHS.

$$2MnO_4^-(aq) + Br^-(aq)$$
$$\longrightarrow MnO_2(s) + BrO_3^-(aq) + H_2O(l) + 2OH^-(aq)$$

It can be seen 1 molecule of H_2O gets cancelled on both sides, we get.

$$2MnO_4^-(aq) + Br^-(aq)$$
$$\longrightarrow MnO_2(s) + BrO_3^-(aq) + H_2O(l) + 2OH^-(aq)$$

is a balanced equation.

96. PbO and PbO_2 react with HCl according to following chemical equations:

$$2PbO + 4HCl \longrightarrow 2PbCl_2 + 2H_2O$$
$$PbO_2 + 4HCl \longrightarrow PbCl_2 + Cl_2 + 2H_2O$$

Why do these compounds differ in their reactivity?

Ans. : $2PbO + 4HCl \longrightarrow 2PbCl_2 + 2H_2O$(Acid base reaction)

$$PbO_2 + 4HCl \longrightarrow PbCl_2 + Cl_2 + 2H_2O(\text{Redox reaction})$$

In reaction (i), O.N. of none of the atoms reaction. It is an acid-base reaction, because PbO is a basic oxide which reacts with HCl acid.undergo a change. Therefore, it is not a redox

The reaction (ii) is a redox reaction in which PbO_2 gets reduced and acts as an oxidizing agent.

97. In neutral or faintly alkaline solution '8' moles of peramanganate anions quantitatively oxidise this sulphate anions to produce x' moles of sulphur containing product. What is magnitude of 'X'.

Ans. : $8MnO_4^- + 3S_2O_3^{2-} + H_2O \longrightarrow 8MnO_2 + 6SO_4^{2-} + 2OH^-$

6 moles of SO_4^{2-} will be formed.

98. i. Identify the oxidant and reductant in the following reactions:

a. $10H^+ + 4Zn(S) + NO_3^-(aq)$
$\longrightarrow 4Zn^{2+}(aq) + NH_4^+(aq) + 3H_2O$

b. $I_2(g) + H_2(g) \longrightarrow 2Hl(g) + S(s)$

ii. Write the anode, cathode and net cell reaction for the following cell:

$Zn(s)|Zn(aq)||Br^-(aq)|Br_2(g), pt$

iii. Give two main functions of salt bridge.

Ans. :

i.

a. Zn is reducing agent because it is losing electrons to form Zn^{2+} i.e. oxidation state is increasing from 0 to +2. NO_3^- is oxidising agent because oxidation state of N is decreasing from +5 to -3, i.e. it is gaining electrons.

b. I is oxidising agent because it is gaining electrons. Its oxidation state is decreasing from 0 to -1 whereas H_2S is reducing agent, the oxidation state of 'S' is increasing from -2 to 0 by losing electrons.

ii.

At anode: $Zn \longrightarrow Zn^{2+} + 2e^-$

At cathode: $Br_2 + 2e^- \longrightarrow 2Br^-$

$Zn + Br_2 \longrightarrow Zn^{2+} + 2Br^-$

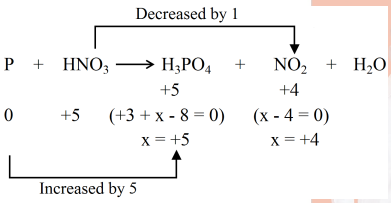
iii.

a. It maintains electroneutrality.

b. It completes internal circuit.

99. Balance $P + HNO_3 \longrightarrow H_3PO_4 + NO_2 + H_2O$ by oxidation number method.

Ans. :



Multiply P by l, HNO_3 by 5, we get $P + 5HNO_3 \longrightarrow H_3PO_4 + 5NO_2 + H_2O$

100. What happens when Cl_2 gas is passed through aqueous solution of KBr? What type of redox reaction is it?

Ans. : $Cl_2g + 2KBr(aq) \longrightarrow 2KCl(aq) + Br_2(l)$ It is non-metal displacement reaction.

101. Balance the following equation:

$Br_2 + H_2O_2 \longrightarrow BrO_3^- + H_2O$ (in acidic medium)

Ans. : $H_2O_2 \longrightarrow H_2O$ (Reduction half reaction)

$H_2O_2 \longrightarrow 2H_2O$ (Balancing oxygen)

$2H^+ + H_2O_2 \longrightarrow 2H_2O$ (Balancing hydrogen)

$2e^- + 2H^+ + H_2O_2 \longrightarrow 2H_2O$ (Balancing charge) ...(i)

$Br_2 \longrightarrow BrO_3^-$ (Oxidation half reaction)

$Br_2 \longrightarrow 2BrO_3^-$ (Balancing bromine)

$6H_2O + Br_2 \longrightarrow 2BrO_3^-$ (Balancing oxygen)

$6H_2O + Br_2 \longrightarrow 2BrO_3^- + 12H^+$ (Balancing hydrogen)

$6H_2O + Br_2 \longrightarrow 2BrO_3^- + 12H^+ + 10e^-$ (Balancing charge) ...(ii)

Multiply equation (i) by 5 and the resultant to equation (ii).

$5H_2O_2 + Br_2 \longrightarrow 2H^+ + 4H_2O + 2BrO_3^-$

102. Calculate the oxidation number of phosphorus in the following species.

a. HPO_3^{2-}

b. PO_4^{3-}

Ans. :

a. Let the oxidation number of phosphorus is x.

$\overset{x}{H}P\overset{2-}{O_3}$

+1 + x + (-2) × 3 = -2

+1 + x - 6 = -2

x - 2 = -2

x = -2 + 5

x = +3

Thus, O.S. of phosphorous is +3.

b. $\overset{x}{P}\ \overset{3-}{O_4}$

x + (-2) × 4 = -3

x - 8 = -3

x = -3 + 8

x = +5

Thus, O.S. of phosphorous in this ion is +5.

103. How many millimoles of potassium dichromate is required to oxidise 24mL of 0.5M Mohr's salt solution in acidic medium?

Ans. : Number of millimoles of $K_2\ Cr_2\ O_7$ present in 24mL of 0.5M solution = 24 × 0.5 = 12

The balanced chemical equation for the redox reaction is

$$K_2Cr_2O_7 + 6(NH_4)_2SO_4 \cdot 6H_2O + 7H_2SO_4 \longrightarrow K_2SO_4 + 6(NH_4)_2SO_4 + 3Fe_2(SO_4)_3 + Cr_2(SO_4)_3 + 43H_2O$$

From the balanced equation, 6 moles Mohr's salt are oxidised by 1 mole of $K_2Cr_2O_7$.

∴ 12 millimoles of Mohr's salt will be oxidised by,

$= \frac{1}{2} \times 12 = 2$ millimoles $K_2\ Cr_2\ O_7$

*** Given Section consists of questions of 3 marks each. [21]**

104. Using the standard electrode potentials given in the Table, predict if the reaction between the following is feasible:

Reaction (Oxidised form + ne⁻ → Reduced form)	$E^{\ominus}$ / V
$F_2(g) + 2e^- \rightarrow 2F^-$	2.87
$Co^{3+} + e^- \rightarrow Co^{2+}$	1.81
$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	1.78
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.51
$Au^{3+} + 3e^- \rightarrow Au(s)$	1.40
$Cl_2(g) + 2e^- \rightarrow 2Cl^-$	1.36
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	1.33
$O_2(g) + 4H^+ + 4e^- \rightarrow 2H_2O$	1.23
$MnO_2(s) + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$	1.23
$Br_2 + 2e^- \rightarrow 2Br^-$	1.09
$NO_3^- + 4H^+ + 3e^- \rightarrow NO(g) + 2H_2O$	0.97
$2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$	0.92
$Ag^+ + e^- \rightarrow Ag(s)$	0.80
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.77
$O_2(g) + 2H^+ + 2e^- \rightarrow H_2O_2$	0.68
$I_2(s) + 2e^- \rightarrow 2I^-$	0.54
$Cu^+ + e^- \rightarrow Cu(s)$	0.52
$Cu^{2+} + 2e^- \rightarrow Cu(s)$	0.34
$AgCl(s) + e^- \rightarrow Ag(s) + Cl^-$	0.22
$AgBr(s) + e^- \rightarrow Ag(s) + Br^-$	0.10
$2H^+ + 2e^- \rightarrow H_2(g)$	**0.00**
$Pb^{2+} + 2e^- \rightarrow Pb(s)$	−0.13
$Sn^{2+} + 2e^- \rightarrow Sn(s)$	−0.14
$Ni^{2+} + 2e^- \rightarrow Ni(s)$	−0.25
$Fe^{2+} + 2e^- \rightarrow Fe(s)$	−0.44
$Cr^{3+} + 3e^- \rightarrow Cr(s)$	−0.74
$Zn^{2+} + 2e^- \rightarrow Zn(s)$	−0.76
$2H_2O + 2e^- \rightarrow H_2(g) + 2OH^-$	−0.83
$Al^{3+} + 3e^- \rightarrow Al(s)$	−1.66
$Mg^{2+} + 2e^- \rightarrow Mg(s)$	−2.36
$Na^+ + e^- \rightarrow Na(s)$	−2.71
$Ca^{2+} + 2e^- \rightarrow Ca(s)$	−2.87
$K^+ + e^- \rightarrow K(s)$	−2.93
$Li^+ + e^- \rightarrow Li(s)$	−3.05

Increasing strength of oxidising agent ↑ / Increasing strength of reducing agent ↓

$Br_2(aq)$ and $Fe^{2+}(aq)$.

Ans. : The possible reaction between $Br_2(aq)$ and $Fe^{2+}(aq)$ is given by,

$$Br_{2(s)} + 2Fe^{2+}_{(aq)} \rightarrow 2Br^-_{(aq)} + 2Fe^{3+}_{(aq)}$$

$$\text{Oxidation half equation:} \quad Fe^{2+}_{(aq)} \longrightarrow Fe^{3+}_{(aq)} + e^-]\times 2 \;; E^\circ = -0.77V$$

$$\text{Reduction half equation:} \quad Br_{2(aq)} + 2e^- \longrightarrow 2Br^-_{(aq)} \;; E^\circ = +1.09V$$

$$Br_{2(aq)} + 2Fe^{2+}_{(aq)} \longrightarrow 2Br^-_{(aq)} + 2Fe^{3+}_{(aq)} \;; E^\circ = -0.32V$$

Here, E° for the overall reaction is positive. Hence, the reaction between $Br_2(aq)$ and $Fe^{2+}(aq)$ is feasible.

105. One mole of N_2H_4 loses 10 moles electrons to form a new compound Y. Assuming that all the nitrogen appears in the new compound, what is the oxidation number of N in Y? There is no change in oxidation state of H.

Ans. : Suppose the oxidation number of N in Y is x

$$(N^{2-})_2 \longrightarrow (2N)^x + 10e$$

$$(\text{as } N_2H_4 \longrightarrow Y + 10e^-)$$

Therefore, 2x - 10 = -4, which geves x = +3.

Hence oxidation number of N in Y = 3.

106. Identify the type of redox reaction taking place in the following.

i. $3\overset{0}{Mg}(s) + \overset{0}{N_2}(g) \longrightarrow \overset{+2\ -3}{Mg_3N_2}(s)$

ii. $\overset{+5\ -2}{V_2O_5}(s) \;\; 5\overset{0}{Ca}(s) \longrightarrow 2\overset{0}{V}(s) + 5\overset{+2\ -2}{CaO}(s)$

iii. $2\overset{+1\ +5\ -2}{KClO_3}(s) \longrightarrow 2\overset{+1\ -1}{KCl}(s) + 3\overset{0}{O_2}(g)$

iv. $\overset{0}{Ca}(s) + 2\overset{+1-2}{H_2Ol} \longrightarrow \overset{+2\ -2+1}{Ca(OH)_2}(aq) + \overset{0}{H_2}(g)$

v. $\overset{0}{Br_2}(l) + 2\overset{-1}{I^-}(aq) \longrightarrow 2\overset{1-}{Br^-}(aq) + \overset{0}{I_2}(s)$

Ans. :

i. Combination reaction.
ii. Displacement reaction.
iii. Decomposition reaction.
iv. Metal displacement reaction.
v. Non-metal displacement reaction.

vi. Disproportionation reaction.

107. How does Cu_2O act as both oxidant and reductant? Explain with proper reactions showing the change of oxidation numbers in each example.

Ans. : Cu^+ undergoes disproportionation to form Cu^{2+} and Cu.

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

Thus, Cu+ or Cu2O acts both as an oxidant as well as reductant.

i. When heated in air, Cu_2O is oxidised to CuO.

$$\overset{+1}{Cu_2}\ \overset{-2}{O} + \frac{1}{2}\ \overset{0}{O_2} \longrightarrow 2\ \overset{+2}{Cu}\ \overset{-2}{O}$$

i.e. Cu_2O acts as a reductant and reduces O_2 to O^{2-}.

ii. When heated with Cu_2S, it oxidises S^{2-} to SO_2 and hence, Cu_2O acts as an oxidant.

$$2\ \overset{+1}{Cu_2}\ \overset{-2}{O} + \overset{+1}{Cu_2}\ \overset{-2}{S} \longrightarrow 6\ \overset{0}{Cu} + \overset{+4}{SO_2}$$

108. Why does fluorine not show disporportionation 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 adeast 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.

109. a. In the following redox reactions, identify the oxidation and reducing agents:

i. $$H_3PO(aq) + 2HgCl_2 + 2H_2O(aq) \longrightarrow H_3PO_4(aq) + 2Hg(l) + 4HCl(aq)$$

ii. $$O_2(g) + PtF_6 g \longrightarrow O_2^+[PtF_6]^{\ominus}(s)$$

b. Why does H_2S acts as reducing agent only whereas SO_2 acts as bot oxidant as wells as rductant?

Ans. :

a.

i. H_3PO_2, is reducing agent, $HgCl_2$ is oxidising agent.

ii. O_2 is reducing agent where as PtF_6 acts as oxidising agent.

b. H_2S has 'S' in -2(lowest) oxidation state, it can only lose electrons acts as reductant. SO_2 has 'S' in +4 oxidation state can show +6 as well as lower oxidation state, therefore, acts as both oxidant as well as reductant.

110. Copper dissolves in dilute nitric acid but not in dilute HCl. Explain.

Ans. : Since, E^0 of Cu^{2+}/ Cu electrode (+0.34V) is higher than that of H^+/ H^2 electrode (0.0V), therefore, H^+ ions cannot oxidise Cu to Cu^{2+} ions and hence, Cu does not dissolve in dil. HCl.

In contrast, the electrode potential of NO_3^- ion, i.e. $\frac{NO_3^-}{NO}$ electrode (+0.97V) is higher than that of copper electrode and hence, it can oxidise Cu to Cu^{2+} ions and hence Cu dissolves in dil. HNO3 due to oxidation of Cu by $\frac{NO_3^-}{NO}$ ions and not by H^+ ions.

*** Given Section consists of questions of 5 marks each.** **[65]**

111. Balance the following redox reactions by ion-electron method:

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

Ans. : Step 1: The two half reactions involved in the given reaction are:

Oxidation half reaction: $\overset{-1}{I}_{(aq)} \rightarrow \overset{0}{I}_{2(s)}$

Reduction half reaction: $\overset{+7}{Mn}O^-_{4(aq)} \rightarrow \overset{+4}{Mn}O_{2(aq)}$

Step 2: Balancing I in the oxidation half reaction, we have:

$$2I^-_{(aq)} \rightarrow I_{2(s)}$$

Now, to balance the charge, we add $2e^-$ to the RHS of the reaction.

$$2I^-_{(aq)} \rightarrow I_{2(s)} + 2e^-$$

Step 3: In the reduction half reaction, the oxidation state of Mn has reduced from +7 to +4. Thus, 3 electrons are added to the LHS of the reaction.

$$MnO^-_{4(aq)} + 3e^- \rightarrow MnO_{2(aq)}$$

Now, to balance the charge, we add 4 OH^- ions to the RHS of the reaction as the reaction is taking place in a basic medium.

$$MnO^-_{4(aq)} + 3e^- \rightarrow MnO_{2(aq)} + 4OH^-$$

Step 4: In this equation, there are 6 O atoms on the RHS and 4 O atoms on the LHS. Therefore, two water molecules are added to the LHS.

$$MnO^-_{4(aq)} + 2H_2O + 3e^- \rightarrow MnO_{2(aq)} + 4OH^-$$

Step 5: Equalising the number of electrons by multiplying the oxidation half reaction by 3 and the reduction half reaction by 2, we have:

$$6I^-_{(aq)} \rightarrow 3I_{2(s)} + 6e^-$$

$$2MnO^-_{4(aq)} + 4H_2O + 6e^- \rightarrow 2MnO_{2(s)} + 8OH^-_{(aq)}$$

Step 6: Adding the two half reactions, we have the net balanced redox reaction as:

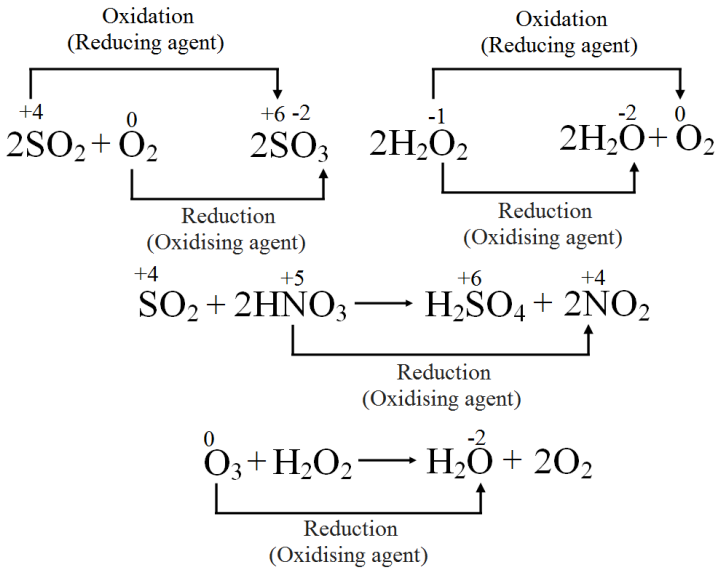
$$6I^-_{(aq)} + 2MnO^-_{4(aq)} + 4H_2O_{(l)} \rightarrow 3I_{2(s)} + 2MnO_{2(s)} + 8OH^-_{(aq)}$$

112. While sulphur dioxide and hydrogen peroxide can act as oxidising as well as reducing agents in their reactions, ozone and nitric acid act only as oxidants. Why?

Ans. :

The oxidation state of sulphur in sulphur dioxide is +4. It can be oxidised to +6 oxidation state or reduced to +2. Therefore, sulphur dioxide acts as a reducing agent as well as oxidising agent. Similarly, the oxidation state of oxygen in hydrogen peroxide is -1. It can be oxidised to O_2 (zero oxidation state) or reduced to H_2O or OH^- (-2 oxidation state) and therefore, acts as reducing as well as oxidising agents.

However, both ozone and nitric acid can only decrease their oxidation number and therefore, act only as oxidising agents.



113. Calculate the oxidation number of sulphur, chromium and nitrogen in H_2SO_5, $Cr_2O_7^{2-}$ and NO_3^-. Suggest structure of these compounds. Count for the fallacy.

Ans. : Oxidation number of sulphur in H_2SO_5:

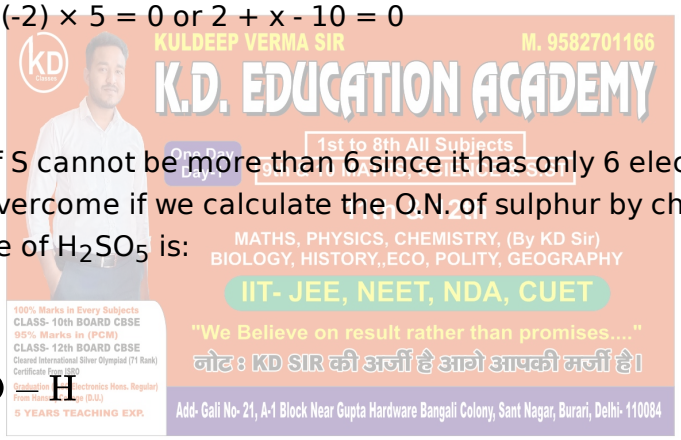
Let the oxidation number of S = x

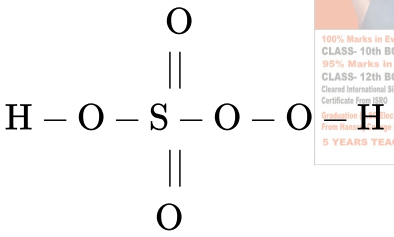
Then, (+1) × 2 + x + (-2) × 5 = 0 or 2 + x - 10 = 0

⇒ x - 8 = 0

∴ x = +8

The maximum O.N. of S cannot be more than 6 since it has only 6 electrons in the valence shell. This fallacy is overcome if we calculate the O.N. of sulphur by chemical bonding method. The structure of H_2SO_5 is:





It has two peroxide oxygen with O.N. = -1

and three oxygens with O.N. = -2

Thus, 2 × (+1) + x + 2(-1) + 3 × (-2) = 0

+2 + x - 2 - 6 = 0 ⇒ x - 6 = 0 ⇒ x = +6

Thus, O.N. of sulphur in H_2SO_5 = +6

Oxidation number of chromium in $Cr_2O_7^{2-}$:

Let the oxidation number of chromium = x

∴ 2x + 7(-2) = -2 ⇒ 2x - 14 = -2

⇒ 2x = -2 + 14

⇒ 2x = +12 ⇒ x = +6

Thus, the oxidation number of chromium = +6

Oxidation number of nitrogen in NO_3^- :

Let the oxidation number of nitrogen = x

Then, x + (-2) × 3 = -1 ⇒ x - 6 = -1

⇒ x = -1 + 6 ⇒ x = +5

Thus, the oxidation number of nitrogen = +5

114. Whenever a reaction between an oxidising agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing agent is in excess and a compound of higher oxidation state is formed if the oxidising agent is in excess. Justify this statement giving three illustrations.

Ans. :

a. C is a reducing agent while O_2 is an oxidising agent. If excess of carbon is burnt in a limited supply of O_2, CO is formed in which the oxidation state of C is +2. If, however, excess of O_2 is used, the initially formed CO gets oxidised to CO_2 in which oxidation state of C is +4.

$$2C(s) + O_2(g) \rightarrow \underset{(\text{Excess})}{\overset{+2}{2CO}}(g); \quad C(s) + O_2(g) \rightarrow \underset{(\text{Excess})}{\overset{+4}{CO_2}}\ (g)$$

b. P_4 is a reducing agent while Cl_2 is an oxidising agent. When excess of P_4 is used, PCl_3 is formed in which the oxidation state of P is +3. If, however, excess of Cl_2 is used, the initially formed PCl_3 reacts further to form PCl_5 in which the oxidation state of P is +5.

$$P_4(s) + 6Cl_2(g) \rightarrow \underset{(\text{Excess})}{\overset{+3}{4PCl_3}}; \quad P_4(s) + 10Cl_2 \rightarrow \underset{(\text{Excess})}{\overset{+5}{4PCl_5}}$$

c. Na is a reducing agent while O_2 is an oxidising agent. When excess of Na is used, sodium oxide is formed in which the oxidation state of O is -2. If, however, excess of O_2 is used, Na_2O_2 is formed in which the oxidation state of O is -1 which is higher than -2.

$$4Na(s) + O_2(g) \rightarrow \underset{(\text{Excess})}{\overset{-2}{Na_2O}}(s); \quad 2Na(s) + 2O_2(g) \rightarrow \underset{(\text{Excess})}{\overset{-1}{Na_2O_2}}(s)$$

115. Write correctly the balanced equations for the following redox reactions using half reactions.

i. $H_2S + Fe^{3+} \rightarrow Fe^{2+} + S + H^+$

ii. $I + IO_3^- + H^+ \rightarrow I_2 + H_2O$

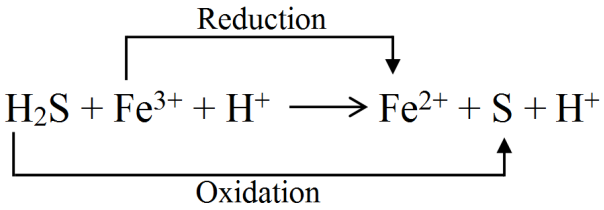
iii. $Bi(s) + NO_3^- + H^+ \rightarrow NO_2 + Bi^{3+} + H_2O$

iv. $I^- + O_2(g) + H_2O \rightarrow I^2 + OH^-$

State what is oxidised to what and what is reduced to what in the reactions expressed by the equations?

Ans. :

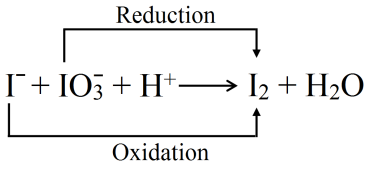
i.



Here, H_2S is oxidised to sulphur (S), and Fe^{3+} is reduced to Fe^{2+}

$2Fe^{3+} + H_2S \rightarrow 2Fe^{2+} + S + 2H^+$

ii.

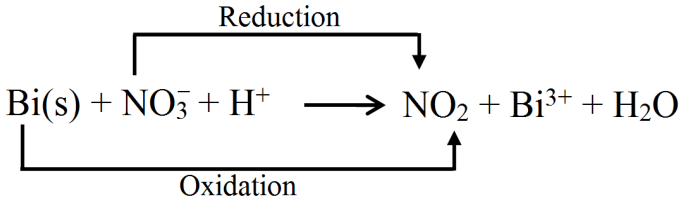


Here, I^- is oxidised to I_2 and IO_3^- is reduced to I^2.

On solving, we get the following balanced equations.

$$2IO^- + 12I^- + 12H^+ \longrightarrow 7I_2 + 6H_2O$$

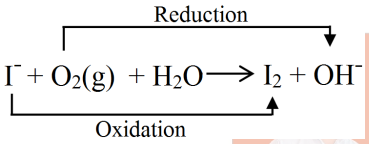
iii.



Here, Bi(s) is oxidised to Bi^{3+}, while NO_3^- is reduced to NO_2.

$$Bi(s) + 3NO_3^- + 6H^+ \longrightarrow 3NO_2 + Bi^{3+} + 3H_2O$$

iv.



Here, I^- is oxidised to I_2 and $O_2(g)$ is reduced to OH^-.

$$4I^- + O_2(g) + 2H_2O \longrightarrow 4OH^- + 2I_2$$

116. i. Use the following reactions to arrange the elements A, B, C and D in order of their redox reactivity:

a. $A + B^+ \rightarrow At + B$

b. $B + D^+ \rightarrow B^+ + D$

c. $C^+ + D \rightarrow$ No reaction

d. $B^+C^+ \rightarrow B^+ + C$

ii. On the basis of above redox activity series, predict which of the following reactions would you expect to occur?

a. $A^+ + C \rightarrow A^+ C^+$

b. $A^+ + D \rightarrow A^+ D^+$

Ans. :

i. The electrochemical series or redox activity is based on the decreasing order of reduction potentials. This means that the species which gets reduced is higher in the electrochemical series as compared to the other which is to be oxidised (lose electrons).

In reaction (a), B^+ gets reduced by A and therefore B is higher than A in electrochemical series.

In reaction (b), D^+ gets reduced by B and therefore, D is higher in the electrochemical series than B.

In reaction (c), C^+ does not get reduced by D, therefore, C is lower than D in electrochemical series. But according to reaction (d), C^+ gets reduced by B and

therefore, C is higher in electrochemical series than B.

Thus, the correct order is,

D > C > B > A

ii. Both reactions do not occur because A cannot be reduced by C as well as D.

117. a. Consider the following redox reaction that produce electricity in a galvanic cell:

i. $2Fe^{3+} + 2Cl^{-} \longrightarrow 2Fe^{2} + Cl_2(g)$

ii. $Cd(s) + I_2 \longrightarrow Cd^{2+} + 2I^{-}$

iii. $2Crs + 3Cu^{2+} \longrightarrow +3Cu(s) + 2Cr^{3+}$

Write the anode and cathode reaction for galvanic cell.

b. Split the following redox reaction into the oxidation and reduction hay reactions:

i. $Zn + Cu^{2+} \longrightarrow 2n^{2+} + Cu$

ii. $Sn^{2} + 2Hg^{+} \longrightarrow Sn^{4+} + Hg_2^{2+}$

Ans. :

a.

i. Anode: $2Cl^{-} \longrightarrow Cl_2 + 2e^{-}$

Cathode: $Fe^{3+} + e^{-} \longrightarrow Fe^{2+}$

ii. Anode: $Cd \longrightarrow Cd^{2} + de^{-}$

Cathode: $I_2 + 2e^{-} \longrightarrow 2I^{-}$

iii. Anode: $Cr \longrightarrow Cr^{3+} + 3e^{-}$

Cathode: $Cu^{2+} \longrightarrow 2e^{-} + Cu$

b.

i. Oxidation half reaction

$Zn \longrightarrow 2n^{2+} + 2e^{-}$

Reduction half reaction

$Cu^{2+} + 2e^{-} \longrightarrow Cu$

ii. Oxidation half reaction

$Sn^{2+} \longrightarrow Sn^{4+} + 2e^{-}$

Reduction half reaction

$Cu^{2+} + 2e^{-} \longrightarrow Cu$

ii. Oxidation half reaction

$Sn^{2+} \longrightarrow Sn^{4+} + 2e^{-}$

$2Hg^{2+} + 2e^{-} \longrightarrow Hg_2^{2+}$

118. Consider the cell reaction of an electrochemical cell:

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

Answer the following questions:

i. Write anode and cathode half reactions.

ii. Mention the direction of flow of electrons.

iii. How is the electroneutrality maintained in solution of two half cells?

iv. Write the formula for calculating standard e.m.f of this cell.

v. How does e.m.f. change when concentration of Ag^{+} is decreased?

Ans. :

i. At anode: $Ni(s) \longrightarrow Ni^{2+}(aq) + 2e^-$

At cathode: $2Ag^+(aq) + 2e^- \longrightarrow 2Ag(s)$

ii. Electrons will flow from nickel to silver i.e. anode to cathode.

iii. Salt bridge contains KCl, K^+ will neutralize negative ion in cathodic half cell and Cl^- will neutralise Ni^{2+} in anodic half cells to maintain electroneutrality.

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

v. E.M.F of cell will decrease when concentration of Ag^+ is decreased.

119. a. Identify the oxidising agent and reducing agent in the following reactions:

i. $MnO_2 + 4HCl \longrightarrow MnCl_2 + Cl_2 + 2H_2O$

ii. $2MnO_4^- 10Cl^- + 16H^+$

$\longrightarrow 2Mn^{2+} + 5Cl_2 + 8H_2O$

b. Calculate the oxidation number of underlined elements in the following speries.

$Pb_3O_4, H_2Cl, PO_4^{3-}$

Ans. :

a.

(i) Increased oxidation number

$MnO_2 + 4HCl \longrightarrow MnCl_2 + Cl_2 + 2H_2O$

Decreased number

(ii) Increased number

$2MnO_4^- + 10HCl + 16H^+ \longrightarrow 2Mn + SCl_2 + 8H_2O$

Decreased number

b.

i. 3x + 4(-2) = 0

⇒ 3x - 8 = 0

$\Rightarrow x = \frac{8}{3}$

ii. x + 2 + (1) + 2(-1) = 0

x + 2 - 2 = 0

⇒ x = 0

iii. x + 4(-2) = 0

⇒ x + 4(-2) = 0

⇒ x - 8 = 0

⇒ x = 8

120. Using electron transfer concept, identify the oxidant and reductant in the following redox reactions.

i. $Zn(s) + 2H^+(aq) \longrightarrow Zn^{2+}(aq) + H_2(g)$

ii. $2[Fe(CN)_6]^{4-}(aq) + H_2O_2(aq) + 2H^+(aq)$

$\longrightarrow 2[Fe(CN)_6]^{3-}(aq) + 2H_2O(l)$

iii. $2[Fe(CN)_6]^{3-}(aq) + 2OH^-(aq) + H_2O_2(aq) \longrightarrow 2[Fe(CN)_6]^{4-}(aq) + O_2(g) + 2H_2(l)$

iv. $BrO_3^-(aq) + F_2(g) + 2OH^-(aq) \longrightarrow BrO_4^-(aq) + F^-(aq) + H_2O(l)$

v. $2NaClO_3(aq) + I_2(aq) \longrightarrow 2NaIO_3(aq) + Cl_2(g)$

Ans. : Oxidants:

i. H^+
ii. H_2O_2
iii. $[Fe(CN)_6]^{3-}$
iv. F_2
v. $NaClO_3$

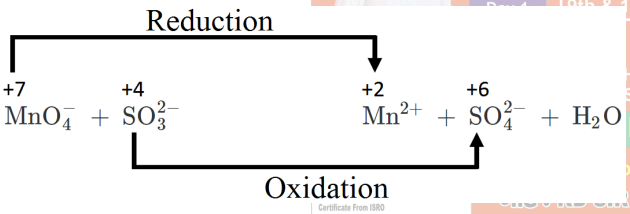
Reductants:

i. Zn
ii. $[Fe(CN)_6]^{4-}$
iii. H_2O_2
iv. BrO_3^-
v. $2I^2$

121. Balance the following ionic equations.

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

Ans. :



Step-1: Separate the equation into two half reactions.

The oxidation number of various atoms are shown below:

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

In this case, chromium undergose reduction, oxidation number decreases from $+6(\text{in } Cr_2O_7^{2-})$ to $+3(\text{in } Cr^{3+})$

Fe^{2+} (O.N. = +2) changes to Fe^{3+} (O.N. = +3). The species undergoing oxidation and reduction are:

Oxidation: $Fe^{2+} \longrightarrow Fe^{3+}$

reduction: $Cr_2O_7^{2-} \longrightarrow Cr^{3+}$

Step-2: Balance each half reaction separately as:

a. $Fe^{2+} \longrightarrow Fe^{3+}$

i. Balance all atoms other than H and O. This step is not needed, because, it is already balanced.
ii. The oxidation number on left is +2 and on right is +3. To account for the difference, the electron is added to the right as: $Fe^{2+} \longrightarrow Fe^{3+} + e^-$
iii. Charge is already balanced.
iv. No need to add H or O.

The balanced half equation is:

$$Fe^{2+} \longrightarrow Fe^{3+} + e^- \ldots (i)$$

Consider the second half equation

$$Cr_2O_7^{2-} \longrightarrow Cr^{3+}$$

i. Balance the atoms other than H and O.

$$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+}$$

ii. The oxidation number of chromium on the left is +6 and on the right is +3. Each chromium atom must gain three electrons. Since there are two Cr atoms, add $6e^-$ on the left.

$$Cr_2O_7^{2-} + 6e^- \longrightarrow 2Cr^{3+}$$

iii. Since the reaction takes place in acidic medium add $14H^+$ on the left to equate the net charge on both sides.

$$Cr_2O_7^{2-} + 6e^- + 14H^+ \longrightarrow 2Cr^{3+}$$

iv. To balance FI atoms, add $7H_2O$ molecules on the right.

$$Cr_2O_7^{2-} + 6e^- + 14H^+ \longrightarrow 2Cr^{3+} + 7H_2O \ldots (ii)$$

This is the balanced half equation.

Step-3: Now add up the two half equations. Multiply eq. (i) by 6 so that electrons are balanced.

$$Fe^{2+} \longrightarrow Fe^{3+} + e^- \times 6$$

$$Cr_2O_7^{2-} + 6e^- + 14H \longrightarrow 2Cr^{3+} + 7H_2O$$

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

122. Write balanced chemical equation for the following reactions:

Permanganate ion (MnO_4^-) reacts with sulphur dioxide gas in acidic medium to produce Mn2+ and hydrogensulphate ion.

(Balance by ion electron method)

Ans. : $2MnO_4^- + 5SO_2 + 2H_2O + H^+ \longrightarrow 5HSO_4^- + 2Mn^{2+}$

Balancing by ion-electron method:

$$\overset{+7}{Mn}O_4^- + \overset{+4}{S}O_2 \longrightarrow \overset{+2}{Mn}^{2+} + \overset{+1}{H}\overset{+6}{S}O_4^- \text{ (Skeletal equation)}$$

$$\text{Oxidation half: } SO_2 \longrightarrow HSO_4^-$$

$$\text{Reduction half : } MnO_4^- \longrightarrow Mn^{2+}$$

$$\text{Oxidation half : } SO_2 \longrightarrow HSO_4^- + 2e^-$$

$$SO_2 + 2H_2O \longrightarrow HSO_4^- + 3H^+ + 2e^- \ldots (i)$$

(Add $2H_2O$ molecules to balance O atoms)

$$MnO_4^- + 5e^- \longrightarrow Mn^{2+}$$

$$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O \ldots (ii)$$

(Add $4H_2O$ molecules to balance O atoms and H atoms)

Add oxidation and reduction half

$$[SO_2 + 2H_2O \longrightarrow HSO_4^- + 3H^+ + 2e^-] \times 5$$

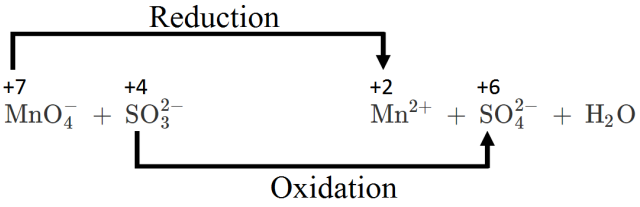
$$[MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O] \times 2$$

$$2MnO_4^- + 5SO_2 + 2H_2O + H^+ \longrightarrow 5HSO_4^- + 2Mn^{2+}$$

123. Balance the following ionic equations.

$$MnO_4^- + SO_3^{2-} + H^+ \longrightarrow Mn^{2+} + SO_4^{2-} + H_2O$$

Ans. :



Dividing the equation into two half reactions:

Oxidation half reaction: $SO_3^{2-} \longrightarrow SO_4^{2-}$

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

Balancing oxidation and reduction half reactions separately as:

Oxidation half reaction

$$SO_3^{2-} \longrightarrow SO_4^{2-}$$

$$SO_3^{2-} \longrightarrow SO_4^{2-} + 2e^-$$

Since the reaction occurs in acidic medium,

$$SO_3^{2-} \longrightarrow SO_4^{2-} + 2e^- + 2H^+$$

$$SO_3^{2-} + H_2O \longrightarrow SO_4^{2-} + 2H^+ + 2e^- \ldots (i)$$

Reduction half reaction

$$MnO_4^- \longrightarrow Mn^{2+}$$

$$MnO_4^- + 5e^- \longrightarrow Mn^{2+}$$

$$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O \ldots (ii)$$

To balance the electrons, multiply eq. (i) by 5 and eq. (ii) by 2 and add

$$2MnO_4^- + 5SO_3^{2-} + 6H^+ \longrightarrow 2Mn^{2+}5SO_4^{2-} + 3H_2O$$

----- Stay away from those people who try to disparage your ambitions. ... -----