PREVIOUS HSE QUESTIONS AND ANSWERS OF THE CHAPTER "REDOX REACTIONS"

1. Balance the following redox equation in acidic medium by half reaction method :

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

   Ans: Step-1: Assign the oxidation number of each element and find out the substance oxidised and reduced.

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

   Here Fe is oxidised and Cr is reduced.

   Step-2: Separate the equation into 2 half reactions -oxidation half reaction and reduction half reaction.

   Oxidation half: $Fe^{2+} \longrightarrow Fe^{3+}$ *Reduction half:* $Cr_2O_7^{2-} \longrightarrow Cr^{3+}$

   Step-3: Balance the atoms other than O and H in each half reaction individually.

   Oxidation half: $Fe^{2+} \longrightarrow Fe^{3+}$ *Reduction half:* $Cr_2O_7^{2-} \longrightarrow 2\ Cr^{3+}$

   Step-4: Now balance O and H atoms. Add H_2O to balance O atoms and H^+ to balance H atoms since the reaction occurs in acidic medium.

   Oxidation half: $Fe^{2+} \longrightarrow Fe^{3+}$ *Reduction half:* $Cr_2O_7^{2-} + 14H^+ \longrightarrow 2\ Cr^{3+} + 7\ H_2O$

   Step -5: Now balance the ionic charges. For this add electrons to one side of the half reaction.

   Oxidation half: $Fe^{2+} \longrightarrow Fe^{3+} + e^-$ *Reduction half:* $Cr_2O_7^{2-} + 14H^+ + 6\ e^- \longrightarrow 2\ Cr^{3+} + 7\ H_2O$

   Step-6: Now add the two half reactions after equating the electrons.

   Oxidation half: $(Fe^{2+} \longrightarrow Fe^{3+} + e^-) \times 6$

   Reduction half: $(Cr_2O_7^{2-} + 14H^+ + 6\ e^- \longrightarrow 2\ Cr^{3+} + 7\ H_2O) \times 1$

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

   Now the equation is balanced.

2. (i) What are disproportionation reactions ? (1)

   (ii) Check whether the reaction $2H_2O_2(l) \longrightarrow 2H_2O(l) + O_2(g)$ is a disproportionation reaction. Justify your answer. (2) [December 2021]

   Ans: (i) It is a type of redox reaction in which an element in one oxidation state is simultaneously oxidised and reduced.

   (ii)

   $2H_2\overset{-1}{O_2}(l) \longrightarrow 2H_2\overset{-2}{O}(l) + \overset{0}{O_2}(g)$

   Here the oxidation number of oxygen is simultaneously increased and decreased. So it is a disproportionation reaction.

3. (i) Oxidation number of oxygen atom in O_2 molecule is _____. (1)

   (ii) In a reaction $2Cu_2O + Cu_2S \longrightarrow 6Cu + SO_2$

   Identify oxidising agent and reducing agent. (2)

   Ans: (i) zero

   (ii) $2\overset{+1}{Cu_2}\overset{-2}{O} + \overset{+1}{Cu_2}\overset{-2}{S} \longrightarrow 6\ \overset{0}{Cu} + \overset{+4}{S}\overset{-2}{O_2}$

   Here the oxidation number of Cu in both Cu_2O and Cu_2S is decreased from +1 to 0. So Cu is reduced and it is the oxidising agent. While the oxidation number of S is increased from -2 to +4. So S in Cu_2S is oxidized and hence it is the reducing agent.

4. (i) Represent the following compounds using stock notation :

   (a) MnO (b) FeO (1)

   (ii) What is oxidation and reduction in terms of oxidation number? (2) [September 2021]

Ans: (i) (a) Mn(II)O   (b) Fe(II)O

(ii) According to oxidation number concept, oxidation is the process of increase in the oxidation number of an element and reduction is the process of decrease in the oxidation number of an element.

5. The oxidation number of oxygen in super oxides is:

   (A) -1   (B) +1   (C) - ½   (D) + ½   (1)

   Ans: - ½

6. Balance the following Redox reaction by oxidation number method or ion-electron method (Acid medium)

   $Cl_2O_7(g) + H_2O_2(aq) \longrightarrow ClO_2^-(aq) + O_2(g) + H^+$ (3)   [December 2020]

   Ans: Step-1: Assign the oxidation number of each element and find out the substance oxidised and reduced.

   $\overset{+7\ -2}{Cl_2O_7} + \overset{+1\ -1}{H_2O_2} \longrightarrow \overset{+3\ -2}{ClO_2^-} + \overset{0}{O_2} + \overset{+1}{H^+}$

   Here the oxidation number of O is increased and that of Cl is decreased. So O in H_2O_2 is oxidised and Cl in Cl_2O_7 is reduced.

   Step-2: Separate the equation into 2 half reactions -oxidation half reaction and reduction half reaction.

   Oxidation half: $H_2O_2 \longrightarrow O_2$ Reduction half: $Cl_2O_7 \longrightarrow ClO_2^-$

   Step-3: Balance the atoms other than O and H in each half reaction individually.

   Oxidation half: $H_2O_2 \longrightarrow O_2$ Reduction half: $Cl_2O_7 \longrightarrow 2\ ClO_2^-$

   Step-4: Now balance O and H atoms. Add H_2O to balance O atoms and H^+ to balance H atoms since the reaction occurs in acidic medium.

   Oxidation half: $H_2O_2 \longrightarrow O_2 + 2\ H^+$ Reduction half: $Cl_2O_7 + 6\ H^+ \longrightarrow 2\ ClO_2^- + 3\ H_2O$

   Step -5: Now balance the ionic charges. For this add electrons to one side of the half reaction.

   Oxidation half: $H_2O_2 \longrightarrow O_2 + 2\ H^+ + 2e^-$ Reduction half: $Cl_2O_7 + 6\ H^+ + 8e^- \longrightarrow 2\ ClO_2^- + 3\ H_2O$

   Step-6: Now add the two half reactions after equating the electrons.

   Oxidation half: $(H_2O_2 \longrightarrow O_2 + 2\ H^+ + 2e^-) \times 4$

   Reduction half: $(Cl_2O_7 + 6\ H^+ + 8e^- \longrightarrow 2\ ClO_2^- + 3\ H_2O) \times 1$

   Overall reaction is: $\mathbf{4\ H_2O_2 + Cl_2O_7 \longrightarrow 4\ O_2 + 2\ H^+ + 2\ ClO_2^- + 3\ H_2O}$

   Now the equation is balanced.

7. The oxidation number of an atom in the elementary form is ……….   (1)

   Ans: zero

8. (a) Justify that the following reaction is a redox reaction

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

   (b) Write the Stock notation of MnO_2.   (1)   [March 2020]

   Ans: (a)

   $\overset{+1\ -2}{H_2S}(g) + \overset{0}{Cl_2}(g) \longrightarrow 2\overset{+1\ -1}{HCl}(g) + \overset{0}{S}(s)$

   Here the oxidation number of sulphur is increased (oxidation) and that of chlorine is decreased (reduction). So it is a redox reaction.

   (b) Mn(IV)O_2

9. (a) In the reaction: $Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \longrightarrow 2\ PbSO_4(s) + 2H_2O(l)$, identify the following. (2)

   (i) The substance oxidised (ii) The substance reduced (iii) The oxidising agent (iv) The reducing agent

   (b) What is disproportionation reaction?   (1)   [July 2019]

Ans: (a) (i) The substance oxidised: Pb

(ii) The substance reduced: Pb in PbO_2

(iii) The oxidising agent: Pb in PbO_2

(iv) The reducing agent: Pb

(b) In a disproportionation reaction, an element in one oxidation state is simultaneously oxidised and reduced.

10. Balance the following Redox process by ion-electron method or oxidation number method :

$$P_4(s) + OH^-(aq) \longrightarrow PH_3(g) + HPO_2^-(aq) \qquad (3) \qquad \text{[March 2019]}$$

Ans: ***Ion-electron method:***

Step-1: Assign the oxidation number of each element and find out the substance oxidised and reduced.

$$\overset{0}{P_4} + OH^- \longrightarrow \overset{-3}{P}H_3 + H\overset{+2}{P}O_2^-$$

Here P_4 is simultaneously oxidised and reduced.

Step-2: Separate the equation into oxidation half reaction and reduction half reaction.

Oxidation half: $P_4 \longrightarrow HPO_2^-$ Reduction half: $P_4 \longrightarrow PH_3$

Step-3: Balance the atoms other than O and H in each half reaction individually.

Oxidation half: $P_4 \longrightarrow 4HPO_2^-$ Reduction half: $P_4 \longrightarrow 4PH_3$

Step-4: Now balance O and H atoms. Add H_2O to balance O atoms and H^+ to balance H atoms. Since the reaction occurs in basic medium also add equal number of OH^- ions on both sides of the equation.

Oxidation half: $P_4 + 8H_2O + 12OH^- \longrightarrow 4HPO_2^- + 12\,H^+ + 12\,OH^-$

Or, $P_4 + 12OH^- \longrightarrow 4HPO_2^- + 4H_2O$

Reduction half: $P_4 + 12\,H^+ + 12\,OH^- \longrightarrow 4PH_3 + 12\,OH^-$

Or, $P_4 + 12\,H_2O \longrightarrow 4PH_3 + 12OH^-$

Step -5: Now balance the ionic charges. For this add electrons to one side of the half reaction.

Oxidation half: $P_4 + 12OH^- \longrightarrow 4HPO_2^- + 4H_2O + 8e^-$

Reduction half: $P_4 + 12\,H_2O + 12e^- \longrightarrow 4PH_3 + 12OH^-$

Step-6: Now add the two half reactions after equating the electrons.

Oxidation half: $(P_4 + 12OH^- \longrightarrow 4HPO_2^- + 4H_2O + 8e^-)x3$

Reduction half: $(P_4 + 12\,H_2O + 12e^- \longrightarrow 4PH_3 + 12OH^-)x2$

Overall reaction is: $5\,P_4 + 12\,OH^- + 12\,H_2O \longrightarrow 12\,HPO_2^- + 8\,PH_3$

<u>OR</u>

Oxidation number method:

***Step 1**: Write the skeletal equation. $P_4 + OH^- \longrightarrow PH_3 + HPO_2^-$*

***Step 2**: Assign the oxidation number of each elements and identify the atoms which undergo change in oxidation number.*

$$\overset{0}{P_4} + \overset{0}{P_4} + OH^- \longrightarrow \overset{-3}{P}H_3 + H\overset{+2}{P}O_2^-$$

***Step 3**: Calculate the change in oxidation number per atom and equate them by multiplying with suitable coefficients.*

O.N decreased by 3

$$\overset{0}{P_4} + \overset{0}{P_4} + OH^- \longrightarrow \overset{-3}{P}H_3 + H\overset{+2}{P}O_2^-$$

O.N increased by 2

$2P_4 + 3P_4 + OH^- \longrightarrow PH_3 + HPO_2^-$

***Step 4**: Balance all the atoms except oxygen and hydrogen.*

$2P_4 + 3P_4 + OH^- \longrightarrow 8\ PH_3 + 12\ HPO_2^-$

***Step 5**: Now equate the ionic charges on both sides. Since the reaction occurs in in basic medium, add 11 OH^- ions on LHS.*

$2P_4 + 3P_4 + OH^- + 11OH^- \longrightarrow 8\ PH_3 + 12\ HPO_2^-$

***Step 6**: Now balance the hydrogen atoms by adding 12 water (H_2O) molecules on LHS.*

$5P_4 + 12OH^- + 12H_2O \longrightarrow 8\ PH_3 + 12\ HPO_2^-$

Now the equation becomes balanced.

11. Redox reactions are classified into four types. Describe any three of them with suitable examples. (3)

[August 2018]

*Ans: **Combination reactions**: A combination reaction may be denoted as A + B → C*

Here either A or B or both A and B must be in the elemental form.

E.g. $C\ (s) + O_2\ (g) \longrightarrow CO_2\ (g)$

***Decomposition reactions:** Decomposition reactions are the opposite of combination reactions. It involves the breakdown of a compound into two or more components, in which at least one must be in the elemental state. It may be denoted as: C → A + B.*

E.g.: $2NH_3\ (g) \longrightarrow N_2\ (g) + 3H_2\ (g)$

***Displacement reactions**: Here an ion (or an atom) in a compound is replaced by an ion (or an atom) of another element. It may be denoted as: X + YZ → XZ + Y*

E.g.: $Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$

12. Balance the following Redox reaction by ion-electron method or oxidation number method (Acid medium)

$Cr_2O_7^{2-}\ (aq) + SO_3^{2-}\ (aq) \longrightarrow Cr^{3+}(aq) + SO_4^{2-}\ (aq)$ (3) [March 2018]

Ans:

13. a) The oxidation number of sulphur in SO_4^{2-} is ……………… a) 3 b) 4 c) 5 d) 6

b) Balance the following equation using oxidation number method.

$Cr_2O_7^{2-}(aq) + SO_3^{2-}(aq) \longrightarrow Cr^{3+}\ (aq) + SO_4^{2-}(aq)$ [In acidic medium] [July 2017]

Ans: (a) 6

(b) Step 1: The skeletal equation is: $Cr_2O_7^{2-} + SO_3^{2-} \rightarrow Cr^{3+} + SO_4^{2-}$

Step 2: Assign oxidation number each element and identify the elements undergoing change in oxidation number.

$$\overset{+6\ -2}{Cr_2O_7^{2-}} + \overset{+4\ -2}{SO_3^{2-}} \rightarrow \overset{+3}{Cr^{3+}} + \overset{+6\ -2}{SO_4^{2-}}$$

Step 3: Calculate the change in oxidation number and make them equal by multiplying with suitable number. Here the oxidation number of Cr is decreased by 3 and that of S is increased by 2. In order to equate them multiply SO_3^{2-} by 3.

$$Cr_2O_7^{2-} + 3\ SO_3^{2-} \rightarrow Cr^{3+} + SO_4^{2-}$$

Step 4: Now balance all the atoms except Oxygen and Hydrogen

$$Cr_2O_7^{2-} + 3\ SO_3^{2-} \rightarrow 2\ Cr^{3+} + 3\ SO_4^{2-}$$

Step 5: Now balance the ionic charges on both sides. Here the net ionic charge on LHS is -8 and on RHS is 0. To equate them add $8H^+$ on LHS, since the reaction takes place in acidic medium.

$$Cr_2O_7^{2-} + 3\ SO_3^{2-} + 8H^+ \rightarrow 2\ Cr^{3+} + 3\ SO_4^{2-}$$

Step 6: Now balance hydrogen atoms by adding sufficient number of H_2O molecules. Here add 4 H_2O molecules on RHS.

$$Cr_2O_7^{2-} + 3\ SO_3^{2-} + 8H^+ \rightarrow 2\ Cr^{3+} + 3\ SO_4^{2-} + 4\ H_2O$$

Now the equation is balanced.

14. Permanganate ion reacts with bromide ion in basic medium to give manganese dioxide and bromated ion. Write the balanced equation for the reaction using oxidation number method. Skeletal equation is:

$MnO_4^- + Br^- \longrightarrow MnO_2 + BrO_3^-$ (3) [March 2017]

Ans: Oxidation number method

Step 1: The skeletal equation is: $MnO_4^- + Br^- \rightarrow MnO_2 + BrO_3^-$

Step 2: Assign oxidation number each element and identify the elements undergoing change in oxidation number.

$$\overset{+7}{Mn}\overset{-2}{O_4^-} + \overset{-1}{Br^-} \rightarrow \overset{+4}{Mn}\overset{-2}{O_2} + \overset{+5}{Br}\overset{-2}{O_3^-}$$

Here the oxidation number of Mn and Br are changed.

Step 3: Calculate the change in oxidation number and make them equal by multiplying with suitable number. Here the oxidation number of Mn is decreased by 3 and that of Br is increased by 6. In order to equate them multiply MnO_4^- by 2.

$$2\ MnO_4^- + Br^- \rightarrow MnO_2 + BrO_3^-$$

Step 4: Now balance all the atoms except Oxygen and Hydrogen

$$2\ MnO_4^- + Br^- \rightarrow 2\ MnO_2 + BrO_3^-$$

Step 5: Now balance the ionic charges on both sides. Here the net ionic charge on LHS is -3 and on RHS is -1. To equate them add 2 OH^- on RHS, since the reaction takes place in basic medium.

$$2\ MnO_4^- + Br^- \rightarrow 2MnO_2 + BrO_3^- + 2OH^-$$

Step 6: Now balance hydrogen atoms by adding sufficient number of H_2O molecules. Here add one H_2O molecule on LHS.

$$2MnO_4^- + Br^- + H_2O \rightarrow 2MnO_2 + BrO_3^- + 2OH^-$$

Now the equation is balanced.

15. In a redox reaction, reduction and oxidation takes place simultaneously.
    a) Write the redox reaction in Daniel cell. (1)
    b) When $CuSO_4$ solution stored in iron vessel, the blue colour changes to pale green. Do you agree with it? Justify. (2) [September 2016]

    Ans: (a) $Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$

    (b) Yes. Iron can displace copper from $CuSO_4$ solution and form $FeSO_4$. So the blue colour changes to pale green.

    $$Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$$

16. Redox reactions can be considered as electron transfer reactions. In an experiment a copper rod is dipped in $AgNO_3$ solution.

a) What happens to the colour of the solution and why? (1)

b) Identify the oxidising and reducing agents in this reaction. (1)

c) Calculate the oxidation number of Cr in $K_2Cr_2O_7$ and P in $H_2P_2O_5$. (1) [March 2016]

Ans: (a) The solution becomes pale blue in colour. This is because Cu displaces Ag from $AgNO_3$ solution. $Cu + 2AgNO_3 \longrightarrow Cu(NO_3)_2 + 2Ag$

(b) Oxidising agent: $AgNO_3$

Reducing agent: Cu

(c) Oxidation no. of Cr in $K_2Cr_2O_7 = +6$

Oxidation no. of P in $H_2P_2O_5 = +4$

17. Identify the oxidant and reductant in the following ionic equation and balance it using oxidation number method.

$$MnO_4^-(aq) + Br^-(aq) + H^+(aq) \longrightarrow Mn^{2+}(aq) + Br_2(l) + H_2O(l)$$ (3) [Sept. 2015]

Ans:

Step 1: The skeletal equation is: $MnO_4^- + Br^- + H^+ \rightarrow Mn^{2+} + Br_2 + H_2O$

Step 2: Assign oxidation number each element and identify the elements undergoing change in oxidation number.

$$\overset{+7}{Mn}\overset{-2}{O_4^-} + \overset{-1}{Br^-} + \overset{+1}{H^+} \rightarrow \overset{+2}{Mn^{2+}} + \overset{0}{Br_2} + \overset{+1}{H_2}\overset{-2}{O}$$

Here the oxidation number of Mn and Br are changed.

Step 3: Calculate the change in oxidation number and make them equal by multiplying with suitable number. Here the oxidation number of Mn is decreased by 5 and that of Br is increased by 1. In order to equate them multiply MnO_4^- by 2 and Br^- by 10 [Since Br is present as Br_2 in RHS]

$$2\,MnO_4^- + 10\,Br^- + H^+ \rightarrow Mn^{2+} + Br_2 + H_2O$$

Step 4: Now balance all the atoms except Oxygen and Hydrogen

$$2\,MnO_4^- + 10\,Br^- + H^+ \rightarrow 2\,Mn^{2+} + 5\,Br_2 + H_2O$$

Step 5: Now balance the ionic charges on both sides. Here the net ionic charge on LHS is -11 and on RHS is +4. To equate them add 15 more H^+ on LHS, since the reaction takes place in acidic medium.

$$2\,MnO_4^- + 10\,Br^- + 16\,H^+ \rightarrow 2\,Mn^{2+} + 5\,Br_2 + H_2O$$

Step6: Now balance hydrogen atoms by adding sufficient number of H_2O molecules. Here add 7 more H_2O molecule on RHS.

$$2\,MnO_4^- + 10\,Br^- + 16\,H^+ \rightarrow 2\,Mn^{2+} + 5\,Br_2 + 8\,H_2O$$

Now the equation is balanced.

18. a) Given the redox reaction:

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

i) Identify the species which undergo reduction and which undergo oxidation.

ii) Identify the reductant and oxidant in the above reaction. (2)

b) Among the following reactions, identify the one which is NOT a redox reaction. (1)

I. $3Mg(s) + N_2(g) \xrightarrow{\Delta} Mg_3N_2(s)$

II. $Fe(s) + 2\,HCl(aq) \longrightarrow FeCl_2(aq) + H_2(g)$

III. $CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$

IV. $2\ NaH(s) \xrightarrow{\Delta} 2Na(s) + H_2(g)$ [March 2015]

Ans: (a) i) Substance oxidised: H_2, Substance reduced: CuO

ii) Reductant: H_2, Oxidant: CuO

(b) $CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$

19. a) Using Stock notation, represent the following compounds: i) $HAuCl_4$ ii) MnO_2 (1)

b) i) Define the electronic concept of oxidation and reduction. (1)

ii) Find out the oxidiser and reducer in the following reaction on the basis of the electronic concept.

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$ (1) [August 2014]

Ans: (a) (i) $HAu(III)Cl_4$ (ii) $Mn(IV)O_2$

(b) (i) According to electronic concept oxidation is the process of removal (losing) of electron and reduction is the process of addition (gaining) of electron.

(ii) Oxidiser: Cl_2 and Reducer: Na

20. a) Write the formula of the following compounds.

i) Nickel (II) sulphate

ii) Tin (IV) oxide (1)

b) Fluorine reacts with ice as given below:

$H_2O(s) + F_2(g) \longrightarrow HF(g) + HOF(g)$ Justify that this is a redox reaction. (2) [March 2014]

Ans: (a) (i) $NiSO_4$ (ii) SnO_2

(b)

$$\overset{-2}{H_2O}(s) + \overset{0}{F_2}(g) \longrightarrow \overset{-1}{HF}(g) + \overset{0}{HOF}(g)$$

Here the oxidation no. of oxygen increases and that of F_2 decreases. So oxygen is oxidised and Fluorine is reduced. Hence it is a redox reaction.

21. a) Calculate the oxidation number of Cr in Cr_2O_3 and S in H_2SO_4. (1)

b) In disproportionation reaction an element in one oxidation state is simultaneously oxidised and reduced. Identify the element undergoing disproportionation in the following reaction:

$P_4 + 3\ OH^- + 3\ H_2O \longrightarrow PH_3 + 3\ H_2PO_2^-$ (2) {September 2013]

Ans: (a) Oxidation no. of Cr in Cr_2O_3 is +3 and the oxidation number of S in H_2SO_4 is +6.

(b)

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

Here the P_4 is simultaneously oxidised and reduced. So it is disproportionate.

22. Competitive electron transfer reactions are utilized in the construction of Galvanic cells.

a) Write the redox reaction involved when metallic cobalt is placed in a nickel sulphate solution. (Note: Only the ionic reaction is required) (1)

b) In the reaction $Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \longrightarrow 2\ PbSO_4(s) + 2\ H_2O(l)$

Identify the following:

i) Substance oxidised ii) Substance reduced iii) Oxidising agent iv) Reducing agent (2) [March 2013]

Ans: (a) $Co + Ni^{2+} \longrightarrow Co^{2+} + Ni$

(b) (i) The substance oxidised: Pb

(ii) The substance reduced: Pb in PbO_2

(iii) The oxidising agent: Pb in PbO_2

(iv) The reducing agent: Pb

23. a) Using stock notation, represent the following compounds - FeO and MnO_2. (1)

b) Redox reactions are those reactions in which oxidation and reduction takes place simultaneously. Write any two redox reactions. (2) [September 2012]

Ans: (a) Fe(II)O and Mn(IV)O_2

(b) $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

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

24. In redox reactions, oxidation and reduction occur simultaneously.

a) How are oxidation and reduction related to the oxidation number? (1)

b) During a group discussion, one of your friends argues that thermal decomposition of $KClO_3$ is a redox reaction while that of $CaCO_3$ is not a redox reaction. Give your opinion and substantiate. (2) [March 2012]

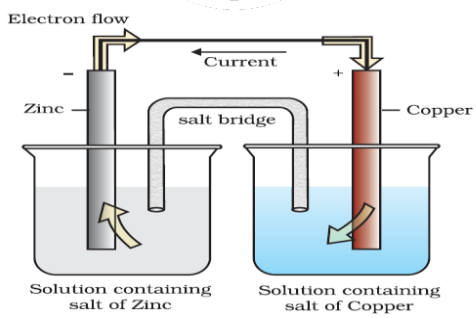
Ans: (a) Oxidation is the process of increase in the oxidation number of an element and reduction is the process of decrease in the oxidation number of an element.

(b) $\overset{+1\ +5\ -2}{2KClO_3} \longrightarrow \overset{+1\ -1}{2KCl} + \overset{0}{3\,O_2}$ *Here Cl is reduced and O is oxidised. This reaction is a redox reaction since there is both oxidation and red...*

$\overset{+2\ +4\ -2}{CaCO_3(s)} \xrightarrow{\Delta} \overset{+2\ -2}{CaO(s)}$ ...

Here there is no change in ox... any species. So it is not a redox reaction.

25. The chemical reactions taking place in e... ...ical cells are redox reactions. A Daniel cell is represented below.



a) As the reaction proceeds in this cell, one of the metal rods gets dissolved in its solution and the other metal gets deposited from the solution to the metal rod. Which metal is getting deposited? (1)

b) Identify the metal which is acting as the oxidising agent in this reaction. (1)

c) Write the chemical equation of the reaction taking place at the first compartment. (1) [October 2011]

Ans: The chemical equation for the reaction is: $Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$

(a) Copper

(b) Cu^{2+}

(c) $Zn \longrightarrow Zn^{2+} + 2e^-$

26. Balance the following equation by the half reaction method.

$Fe^{2+}(aq) + Cr_2O_7^{2-}(aq) + H^+(aq) \longrightarrow Fe^{3+}(aq) + Cr^{3+}(aq) + H_2O(l)$ (3) [March 2011]

Ans: Refer the answer of the question number 1

27. A farmer prepared 1% solution of copper sulphate using iron rod as the stirrer for preparing Bordeaux mixture. Next day he noticed that the blue colour almost disappeared and the iron rod get coated with reddish brown material.
    a) What is the reddish brown material deposited on the iron rod? (1)
    b) Account for the colour change of the solution. (1)
    c) Justify the above phenomenon as a redox reaction. (1) [September 2010]

    Ans: (a) Copper

    b) Here iron displaces copper from $CuSO_4$ solution and form $FeSO_4$. So the blue colour disappears.

    c) $\overset{0}{Fe} + \overset{+2}{Cu}\overset{+6}{S}\overset{-2}{O_4} \longrightarrow \overset{+2}{Fe}\overset{-6}{S}\overset{-2}{O_4} + \overset{0}{Cu}$

    Here the oxidation number of Fe increases and hence it is oxidised, while that of Cu decreases, so it is reduced. Since there is both oxidation and reduction, it is a redox reaction.

28. Chemical reactions which involve oxidation and reduction are called redox reactions. The unbalanced equation in the ionic form of a redox reaction is shown below.

    $Fe^{2+}(aq) + Cr_2O_7^{2-}(aq) \xrightarrow{\text{acidic medium}} Fe^{3+}(aq) + Cr^{3+}(aq)$

    a) Identify the oxidising agent in this reaction. (1)
    b) Name the species getting oxidized in the above reaction. (1)
    c) Balance the above equation by oxidation number method. (3) [March 2010]

Ans: (a) Oxidising agent – Cr in $Cr_2O_7^{2-}$

(b) Fe^{2+}

(c) Oxidation number method

Step 1: The skeletal equation is: $Fe^{2+} + Cr_2O_7^{2-} \rightarrow Fe^{3+} + Cr^{3+}$

Step 2: Assign oxidation number each element and identify the elements undergoing change in oxidation number.

$$\overset{+2}{Fe^{2+}} + \overset{+6}{Cr_2}\overset{-2}{O_7^{2-}} \rightarrow \overset{+3}{Fe^{3+}} + \overset{+3}{Cr^{3+}}$$

Step 3: Calculate the change in oxidation number and make them equal by multiplying with suitable number. Here the oxidation number of Cr is decreased by 3 and that of Fe is increased by 1. In order to equate them multiply Fe^{2+} by 6.

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

Step 4: Now balance all the atoms except Oxygen and Hydrogen

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

Step 5: Now balance the ionic charges on both sides. Here the net ionic charge on LHS is +11 and on RHS is +24. To equate them add 13 more H^+ on LHS, since the reaction takes place in acidic medium.

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

Step 6: Now balance hydrogen atoms by adding sufficient number of H_2O molecules. Here add 7 H_2O molecules on RHS.

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

Now the equation is balanced.

29. Fill in the blanks.

a) The oxidation state of Cl in $HClO_4$ is ……….. (1)

b) A reducing agent is a substance which ……………… electrons in a chemical reaction. (1)

c) Among the elements Fluorine and Iodine, ………………… exhibit both positive and negative oxidation states. (1) [March 2009]

Ans: (a) +7

(b) donates or loses

(c) Iodine

30. a) Both HCl and NaH contain H, but the oxidation states of H in them are different. What is the oxidation state of H in each compound? (2)

b) What is the oxidation state of 'S' in SO_4^{2-}? (1) [June 2008]

Ans: (a) Oxidation state of H in HCl is +1 and in NaH is -1.

(b) +6

31. a) A compound is formed between oxygen and fluorine. Do you know whether it is oxygen fluoride or fluorine oxide? Explain. (2)

b) NO and HNO_3 are two compounds of nitrogen. In which of them N is more oxidised? (1) [February 2008]

Ans: (a) Oxygen fluoride (OF_2), since F is more electronegative than Oxygen.

(b) In HNO_3, since here the oxidation number of N is +5.
