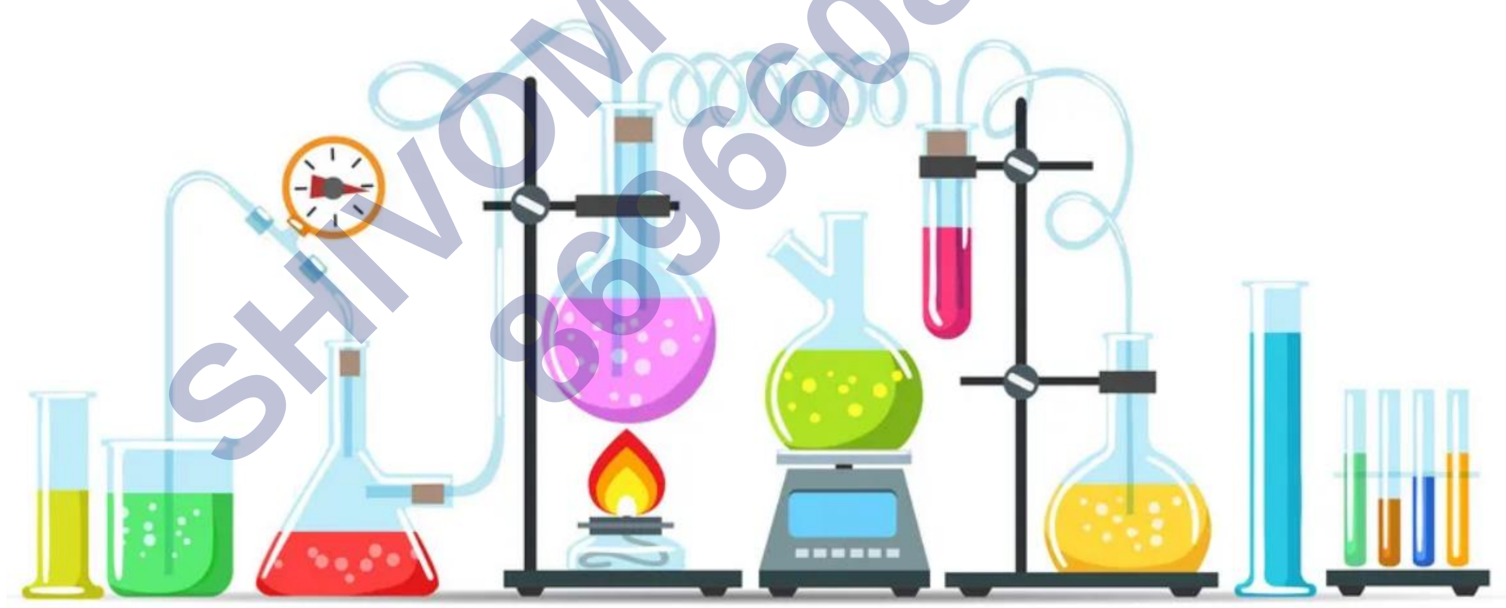
CHEMISTRY

CHAPTER 8: REDOX REACTIONS



REDOX REACTIONS

Introduction

Redox reaction is related to gain or loss of electrons. Reaction in which oxidation and reduction takes place simultaneously is called redox reaction. This chapter deals with problems based on redox reactions, oxidation number and balancing of redox reactions by ion, electron method and oxidation number method.

Oxidation Reactions

Oxidation is defined as the addition of oxygen/electronegative element to a substance or rememoval of hydrogen/ electropositive element from a susbtance.

$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$

$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$

Reduction Reactions

Reduction is defined as the memoval of oxygen/electronegative element from a substance or addition of hydrogen or electropositive element to a substance.

$2FeCl_3(aq) + H_2(g) \longrightarrow 2FeCl_2(aq) + 2HCl(aq)$

$2HgO(s) \longrightarrow 2Hg(l) + O_2(g)$

Oxidation Number or Oxidation State

Oxidation number for an element is the arbitrary charge present on one atom when all other atoms bonded to it are removed. For example, if we consider a molecule of HCl, the Cl atom is more electronegative than H-atom, therefore, the bonded electrons will go with more electronegative chlorine atom resulting in formation of H^+ and Cl^- ions. So oxidation number of H and Cl in HCl are +1 and –1 respectively.

The following points are important to determine the oxidation number of an element.

1. The oxidation number of an atom in pure elemental form is considered to be zero. e.g., H_2, O_2, Na, Mg.

2. Oxidation number of any element in simple monoatomic ion will be equal to the charge on that ion, for example, oxidation number of Na in Na^{+} is +1.
3. Oxidation number of fluorine in its compound with other elements is always -1.
4. Oxidation number of oxygen is generally -2 but in case of peroxide (H_2O_2) oxygen has oxidation number -1. In a compound OF_2 the oxidation number of oxygen is +2.
5. The oxidation number of alkali metals (Na, K) and alkaline earth metals (Ca, Mg) are +1 and +2 respectively.
6. The oxidation number of halogens is generally -1 when they are bonded to less electronegative elements.
7. Oxidation number of hydrogen is generally +1 in most of its compounds but in case of metal hydride (NaH, CaH_2) the oxidation number is hydrogen is -1.
8. The algebraic sum of the oxidation numbers of all the atoms in a neutral compound is zero. In an ion, the algebraic sum of oxidation number is equal to the charge on that ion.

Oxidising and Reducing Agent

A substance which undergoes oxidation acts as a reducing agent while a substance which undergoes reduction acts as an oxidising agent. For example, we take a redox reaction,

$Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$

In this reaction, Zn is oxidised to Zn^{2+} so Zn is reducing agent and Cu^{2+} is reduced to Cu so Cu^{2+} is an oxidising agent.

Types of Redox Reactions

1. **Combination reactions**

A combination reaction is a reaction in which two or more substances combine to form a single new substance. Combination reactions can also be called synthesis reactions. The general form of a combination reaction is:

$A + B \longrightarrow AB$

$Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$

2. **Decomposition reactions**

A decomposition reaction is a reaction in which a compound breaks down into two or more simpler substances. The general form of a decomposition reaction is:

$AB \longrightarrow A + B$

$2HgO(s) \longrightarrow 2Hg(l) + O_2(g)$

3. **Displacement reactions**

Displacement reaction is a chemical reaction in which a more reactive element displaces a less reactive element from its compound.

$CuSO_4(aq) + Zn(s) \longrightarrow ZnSO_4(aq) + Cu$

4. **Disproportionation reactions**

The reactions in which a single reactant is oxidized and reduced is known as Disproportionation reactions. The disproportionation reaction is given below.

$2H_2O_2 \longrightarrow 2H_2O + O_2$

Balancing of Redox Reactions

a. Oxidation Number Method

In this method number of electrons lost in oxidation must be equal to number of electrons gained in reduction. Following rules are followed for balancing of reactions:

1. Write the skeletal equation of all the reactants and products of the reaction.
2. Indicate the oxidation number of each element and identify the elements undergoing change in oxidation number.
3. Equalize the increase or decrease in oxidation number by multiplying both reactants and products undergoing change in oxidation number by a suitable integer.
4. Balance all atoms other than H and O, then balance O atom by adding water molecules to the side short of O-atoms.
5. In case of ionic reactions:

   i. **For acidic medium:** First balance O atoms by adding H_2O molecules to the side deficient in O atoms and then balance H-atoms by adding H+ ions to the side deficient in H atoms.

   ii. **For basic medium:** First balance O atoms by adding H_2O molecules to whatever side deficient in O atoms. The H atoms are then balanced by adding H_2O molecules equal in number to the deficiency of H atoms and an equal number of OH– ions are added to the opposite side of the equations.

Balance the ionic equation in alkaline medium

$$Cr(OH)_3 + IO_3^- \longrightarrow I^- + CrO_4^{2-}$$

Decreases by 6

$$\overset{+3}{Cr}(OH)_3 + \overset{+5}{I}O_3^- \longrightarrow I^- + \overset{+6}{Cr}O_4^{2-}$$

Increases by 3

$$2Cr(OH)_3 + IO_3^- \longrightarrow I^- + 2CrO_4^{2-}$$

Balancing of O and H

$$2Cr(OH)_3 + IO_3^- + 4OH^- \longrightarrow I^- + 2CrO_4^{2-} + H_2O + 4H_2O$$

Now, $2Cr(OH)_3 + IO_3^- + 4OH^- \longrightarrow I^- + 2CrO_4^{2-} + 5H_2O$

b. Ion-Electron Method

1. Write the skeleton equation and indicate the oxidation number of all the elements which appear in the skeletal equation above their respective symbols.

2. Find out the species which are oxidised and which are reduced.

3. Split the skeleton equation into two half reactions, i.e., oxidation half reaction and reduction half reaction.

4. Balance the two half reaction equations separately by the rules described below:

   i. In each half reaction, 1st balance the atoms of the elements which have undergone a change in oxidation number.

   ii. Add electrons to whatever side is necessary to make up the difference in oxidation number in each half reaction.

   iii. Balance oxygen atoms by adding required number of H_2O molecules to the side deficient in O atoms.

   iv. In the acidic medium, H atoms are balanced by adding H^+ ions to the side deficient in H-atoms. However, in the basic medium, H atoms are balanced by adding H_2O molecules equal in number to the deficiency of H atoms and an equal number OH^- ions are included in the opposite side of the equation.

5. The two half reactions are then multiplied by suitable integers so that the total number of electrons gained in one half of the reaction is equal to the number of electrons lost in the other

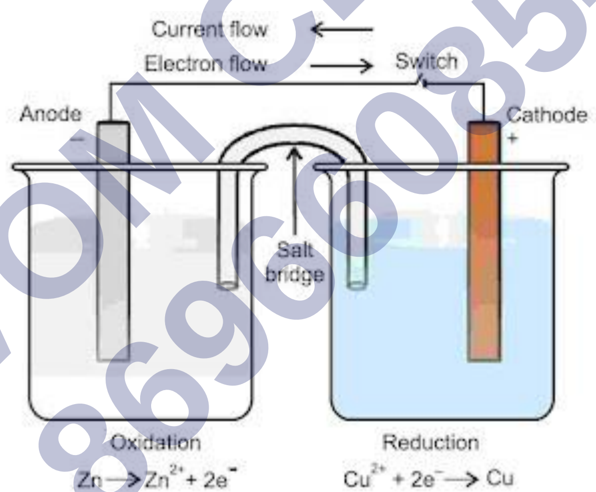
half reaction. The two half reactions are then added up.

6. To verify whether the equation thus obtained is balanced or not, the total charge on either side of the equation must be equal.

Galvanic Cell and Electrode Potential

A galvanic cell or voltaic cell is simple electrochemical cell in which a redox reaction is used to convert chemical energy into electrical energy. It means electricity can be generated with the help of redox reaction in which oxidation and reduction takes place in two separate compartments. Each compartment consists of a metallic conductor and dipped in suitable electrolytic solution of same metal. Metallic rod acts as electrode.

The compartment having electrode dipped in solution of electrolyte is known as half-cell and a half cell has a redox couple. A redox couple means a solution having reduced and oxidised form of a substance together, taking part in oxidation or reduction half reaction. It is depicted as M^{+n} / M i.e., oxidised form / reduced form. To prepare a galvanic cell two half cells are externally connected through a conducting wire and internally through salt bridge.



Anodic oxidation: $Zn_2 \longrightarrow Zn^{+2}(aq) + 2e(s)$

Cathodic reduction: $Cu^{+2}(aq) + 2e \longrightarrow Cu(s)$

Net reaction: $Zn(s) + Cu^{+2}(aq) \longrightarrow Zn^{+2}(aq) + Cu(s)$

This cell can be briefly presented in one line, known as cell notation i.e.,

$Zn \mid Zn^{+2} \mid\mid Cu^{+2} \mid Cu$

Summary-

1. **Oxidation number:** Charge on atom which appears on it when it is present in the combined state.
2. Sum of the oxidation states in a compound/ion should be equal to the zero or to the net charge on the ion.
3. Some elements show variable oxidation states.
4. **Oxidation:** The process in which electrons are lost.
5. **Reduction:** The process in which electrons are gained.
6. **Oxidising agent:** A substance which oxidises the other.
7. **Reducing agent:** A substance which reduces the other.
8. **Redox reaction:** When oxidation and reduction take place together is known as redox reaction.
9. **Disproportionation reaction:** The reaction in which same species is simultaneously oxidised as well as reduced.
10. The change in oxidation state of any element in a compound is useful in calculating the equivalent weight.
11. **Electrochemical series:** Arrangement of E°red of different electrodes in increasing order of electrode potential.
12. **Electrode Potential:** The tendency of an electrode to lose or gain electrons is called electrode potential.
13. The standard electrode potentials of a large number of electrodes have been determined using standard hydrogen electrode as the reference electrode. By convention, the standard electrode potential (E°) of hydrogen electrode is 0.00 volts.

Class : 11th Chemistry
Chapter- 8: Redox Reactions

Redox Reactions

Balancing of Redox Reactions

Two Methods

O.N. Method

Step 1: Write the correct formula for each reactant and product.
Step 2: Identify atoms undergoing change in O.N.
Step 3: Calculate increase/decrease in O.N per atom and for entire ion or molecule. If unequal, multiply by suitable number to make equal.
Step 4: Add H^+/OH^- ion to make total ionic charges of reactants and product equal.
Step 5: Equalize H^+ on two sides by adding water.

Half-Reaction Method

Step 1: Produce, unbalanced equation for the reaction in ionic form.
Step 2: Separate equation into two half-reactions.
Step 3: Individually balance atoms other than O and H.
Step 4: For reactions occurring in acidic medium, add H_2O to balance O atoms and H^+ to balance H atoms.
Step 5: Balance charges by adding electrons to one side of the half reaction.
Step 6: Add two half reactions and cancel the electrons on each side.
Step 7: Verify if equation has same type and number of atoms and same charges on both sides of the equation.

Oxidation Number

Oxidation state of an element in a compound ascertained according to a set of rules formulated on the basis that pair in a covalent bond belongs to electronegative element.

Rules for O.N.

1. In elements, each atom bears an oxidation number of zero.
2. For ions with only one atom, O.N. is equal to the charge on ion.
3. O.N. of oxygen in most compounds is -2.
4. O.N. of hydrogen is +1, except when it is bonded to metals in binary compound.
5. Halogens have on O.N. of -1, when they occur as halide ions in their compound.
6. Algebraic sum of O.N. of all the atoms in a compound must be zero.

Competitive Electron Transfer Reactions

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

Release of $2e^-$ (Zn → Zn^{2+}); Gain of $2e^-$ (Cu^{2+} → Cu)

Oxidation: Addition of oxygen to an element/ compound or removal of hydrogen from a substance

Loss of electrons → **Reducing agent:** Donor of electron(s) → **Lowers O.N.**

Reduction: Removal of oxygen from a substance or addition of hydrogen to a substance

Gain of electrons → **Oxidising agent:** Acceptor of electron(s) → **Increases O.N.**

Application

Study of electrode processes and cells.

Redox Reactions and Electrode Processes

Redox Couple: Having together oxidised and reduced forms of a substance involved in oxidation/reduction half reaction.
Electrode Potential: Potential associated with each electrode.
Standard Electrode Potential: Electrode potential of concentration of each species is unity and carried out at 298K.
Negative E: Stronger reducing agent than H^+/H_2 couple.
Positive E: Weaker reducing agent than couple. H^+/H_2

Types of Redox Reactions

1. Combination reactions: A + B → C; Either A and B or both A and B must be in elemental form.
2. Decomposition reactions: Leads to the breakdown of a compound into two or more components at least one of which must be in the elemental state $CuSO_4(aq) + Zn(s) \rightarrow Cu(s) + ZnSO_4(aq)$
3. Displacement reactions: An ion/ atom in a compound, is replaced by an ion/atom of another element, X + YZ → XZ + Y
(a) Metal displacement: $CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$
(b) Non-Metal displacement: Includes hydrogen displacement.
4. Disproportionation reaction: An element in one oxidation state is simultaneously oxidised and reduced.

Important Questions

Multiple Choice questions-

Question 1. $KMnO_4$ reacts with oxalic acid according to the equation $2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$ Here 20 mL of 0.1 M $KMnO_4$ is equivalent to

(a) 50 mL of 0.5 M $C_2H_2O_4$

(b) 20 mL of 0.1 M $C_2H_2O_4$

(c) 20 mL of 0.5 M $C_2H_2O_4$

(d) 50 mL of 0.1 M $C_2H_2O_4$

Question 2. Which of the following is a redox reaction?

(a) $NaCl + KNO_3 \rightarrow NaNO_3 + KCl$

(b) $Mg(OH)_2 + 2NH_4Cl \rightarrow MgCl_2 + 2NH_4OH$

(c) $CaC_2O_4 + 2HCl \rightarrow CaCl_2 + H_2C_2O_4$

(d) $2Zn + 2AgCN \rightarrow 2Ag + Zn(CN)_2$

Question 3. The reduction potential values of M, N and O are +2.46 V, -1.13 V, -3.13 V respectively. Which of the following orders is correct regarding their reducing property?

(a) O > N > M

(b) M > O > N

(c) M > N > O

(d) O > M > N

Question 4. Which of the following processes does not involve either oxidation or reduction?

(a) Formation of slaked lime from quick lime

(b) Heating Mercuric Oxide

(c) Formation of Manganese Chloride from Manganese oxide

(d) Formation of Zinc from Zinc blende

Question 5. The number of moles of $KMnO_4$ reduced by one mole of KI in alkaline medium is

(a) One

(b) Two

(c) Five

(d) One fifth.

Question 6. What is known as Autooxidation?

(a) Formation of H_2O by the oxidation of H_2O_2.

(b) Formation of H_2O_2 by the oxidation of H_2O.

(c) Both (1) and (2) are true

(d) None of the above

Question 7. Which of the following statements regarding sulphur is incorrect?

(a) S_2 molecule is paramagnetic.

(b) The vapour at 200° C consists mostly of S_8 rings.

(c) At 600°C the gas mainly consists of S_2 molecules.

(d) The oxidation state of Sulphur is never less than +4 in its compounds.

Question 8. The oxidation number of Xe in $BaXeO_6$ is

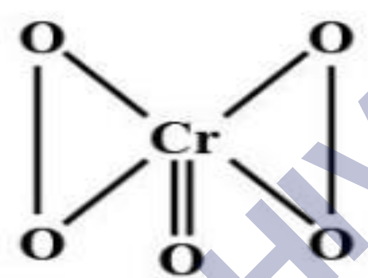
(a) 8

(b) 6

(c) 4

(d) 10

Question 9. CrO_5 has structure as shown, the oxidation number of chromium in the compound is?



(a) +10

(b) +6

(c) +4

(d) +5

Question 10. Pure water is bad conductor of electricity because

(a) It has high boiling point

(b) It is almost unionised

(c) Its molecules are associated with H- bonds

(d) Its pH is 7 at 25°C

Question 11. The oxidation process involves

(a) Increase in oxidation number

(b) Decrease in oxidation number

(c) No change in oxidation number

(d) None of the above

Question 12. The ionic mobility of alkali metal ions in aqueous solution is maximum for

(a) Li+

(b) Na+

(c) K+

(d) Rb+

Question 13. Pure water is bad conductor of electricity because

(a) It has high boiling point

(b) It is almost unionised

(c) Its molecules are associated with H- bonds

(d) Its pH is 7 at 25°C

Question 14. The oxidation number of Fe in $K_4[Fe(CN)_6]$ is

(a) 3

(b) 4

(c) 2

(d) Zero

Question 15. A standard hydrogen electrode has zero electrode potential because

(a) Hydrogen is easiest to oxidise

(b) This electrode potential is assumed to be zero

(c) Hydrogen atom has only one electron

(d) Hydrogen is the lightest element

Very Short:

1. What are redox reactions? Give an example.
2. Define oxidation and reduction in terms of electrons.
3. Define an oxidizing agent. Name the best oxidizing agent.
4. What is meant by reducing agent? Name the best reducing agent.
5. In the reaction $MnO_2 + 4HCl \rightarrow MnCl_2 + H_2O$ which species is oxidized?

6. What is the oxidation state of Ni in Ni $(CO)_4$?
7. What is a redox couple?
8. Define oxidation and reduction in the term of oxidation numbers.
9. What is the sum of oxidation numbers of all atoms in $HIO_4$4?
10. What is the oxidation number of N in $(NH_4)_2$ SO_2?

Short Questions:

1. HNO3 acts only as an oxidant whereas HNOz acts both as an oxidant and reductant. Why?
2. Balance the following equation by the ion-electron method.

   Zn(s) + NO3– → Zn2+(aq) + NH+ (aq) + H2O(l) (In acid solution)
3. Balance the following equation in acidic medium by oxidation number method.
4. Indicate the oxidising and reducing agent in the following reactions:
5. Which of the following redox reaction is oxidation & which is reduction?
6. What are the minimum and maximum oxidation numbers shown by sulfur?

Long Questions:

1. What are the minimum and maximum oxidation numbers shown by sulfur?
2. Starting with the correctly balanced half-reaction, write the overall net ionic equation for the following change:
3. Write the method used for balancing redox reaction by oxidation number method.
4. Determine the oxidation number of O in the following: OF2, Na2O2 & CH3COOH

   (i) OF_2

   (ii) Na_2O_2

   (ii) CH_3COOH
5. Determine the volume 6f M/8 $KMnO_4$ solution required to react completely with 25.0 cm3 of M/4 $FeSO_4$ solution in an acidic medium.

Assertion Reason Questions:

1. In the following questions, a statement of Assertion (A) followed by a statement of Reason (R) is given. Choose the correct option out of the choices given below each question.

   Assertion (A) : Among halogens fluorine is the best oxidant.

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

(i) Both A and R are true and R is the correct explanation of A.

(ii) Both A and R are true but R is not the correct explanation of A.

(iii) A is true but R is false.

(iv) Both A and R are false.

2. In the following questions, a statement of Assertion (A) followed by a statement of Reason (R) is given. Choose the correct option out of the choices given below each question.

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

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

(i) Both A and R are true and R is the correct explanation of A.

(ii) Both A and R are true but R is not the correct explanation of A.

(iii) A is true but R is false.

(iv) Both A and R are false.

Case Study Based Question:

1. Read the passage given below and answer the following questions:

The oxidation state of an individual atom is 0. The total oxidation state of all atoms in a neutral species is 0 and in an ion is equal to the ion charge. Group 1 metals have an oxidation state of + 1 and group 2 an oxidation state of + 2.

The oxidation state of fluorine is − 1 in compounds. Hydrogen generally has an oxidation state of + 1in compounds. Oxygen generally has an oxidation state of − 2 in compounds.

In binary metal compounds, group 17 elements have an oxidation state of − 1, group 16 elements of − 2, and group 15 elements of − 3. The sum of the oxidation states is equal to zero for neutral compounds and equal to the charge for polyatomic ion species. An atom is oxidised if its oxidation number increases and an atom is reduced if its oxidation number decreases.

The atom that is oxidised is the reducing agent and the atom that is reduced is the oxidising agent.

(1) One mole of acidified $K_2 Cr_2 O_7$ on reaction with excess KI will liberate n mole of I_2 then the value of n is:

(a) 6

(b) 1

(c) 3

(d) 7

(2) When electrons are transferred from Zn to Cu^{2+} in copper sulphate solution, the energy (heat) is:

(a) Absorbed

(b) Evolved

(c) Consumed

(d) Both (a) and (b)

(3) Negative $E^\ominus$ indicates that redox couple is

(a) Weaker reducing agent than H^+/H_2 couple

(b) Stronger reducing agent than H^+/H_2 couple

(c) Stronger oxidising agent than H^+/H_2 couple

(d) Weaker oxidising agent than H^+/H_2 couple

(4) Which of the following statements is/are incorrect?

(a) The reactants, which undergo oxidation and reduction are called reductant and oxidant respectively

(b) In redox reaction, the oxidation number of oxidant increases, while that of reductant decreases

(c) HNO_2 acts as an oxidising as well as reducing agent

(d) Oxidation is the process, in which electrons are lost

2. Read the passage given below and answer the following questions:

The concept of electron transfer is found unable to explain the redox changes or electron shift in case of covalent compounds.

To explain these changes a new concept, called oxidation number is introduced. Oxidation number is defined as the charge that an atom of the element has in its ion or appear to have when present in the combined state with other atoms. In other words, it is also defined as the charge that an atom appear to have in a compound when all other atoms are removed as ions from the compound.

The following steps are involved while calculating the oxidation number of an atom in a given compound/ ion.

Step I Write down the formula of given compound/ion leaving some space between the atoms.

Step II Write the oxidation state of each element above its atoms. Write down x above the atom, oxidation state of which we have to find out.

Step III Multiply the oxidation numbers of each element with the number of atoms of that element present in the compound. Enclose the product in a bracket.

Step IV Equate the algebraic sum of the oxidation numbers of all the atoms present in compound to zero or to the charge in case of ionic species charge on the ion.

Step V Solve the equation obtained for the value of x.

(1) Highest oxidation state of Mn is present in:

(a)$KMnO_4$

(b)K_2MnO_4

(c)M_2nO_3

(d)MnO_2

(2) Identify the element which never has positive oxidation number in any of its compound?

(a) Oxygen

(b) Chlorine

(c) Fluorine

(d) Bromine

(3) When a manganous salt is fused with a mixture of KNO_3 and solid NaOH, the oxidation number of Mn changes, from + 2 to:

(a) + 4

(b) + 3

(c) + 6

(d) + 7

(4) The brown ring complex compound is formulated as $[Fe(H_2O)_5 NO]SO_4$. What will be the oxidation state of iron in the given complex?

(a) + 2

(b) + 3

(c) + 4

(d) + 1

Answer Key:

MCQ

1. (d) 50 mL of 0.1 M$C_2H_2O_4$
2. (d) $2Zn + 2AgCN \rightarrow 2Ag + Zn(CN)_2$
3. (d) O > M > N
4. (a) Formation of slaked lime from quick lime
5. (b) Two
6. (b) Formation of H_2O_2 by the oxidation of H_2O.
7. (d) The oxidation state of sulphur is never less than +4 in its compounds.
8. (d) 10
9. (b) +6
10. (b) It is almost unionised
11. (a) Increase in oxidation number
12. (d) Rb^+
13. (b) It is almost unionised
14. (c) 2
15. (b) This electrode potential is assumed to be zero

Very Short Answer:

1. Redox reaction is a reaction in which oxidation and reduction take place simultaneously, e.g.
$Zn(s) + Cu^{2}+(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

2. Oxidation involves loss and reduction involves the gain of electrons.

3. The oxidizing agent is a substance that can gain electrons easily. F_2 is the best oxidizing agent.

4. The reducing agent is a substance that can lose electrons easily. Li is the best reducing agent

(ii)Dissociation increases, i.e., the equilibrium shifts forward.

5. HCl is oxidized to Cl_2

6. Zero

7. The redox couple consists of the oxidized and reduced form of the same substance taking part in an oxidation or reduction half-reaction, for example.

$Zn^{2+}(aq) / Zn$, $Cl_2 / Cl^-(aq)$ etc.

8. Oxidation involves an increase in oxidation number while reduction involves a decrease in oxidation number.

$Sn^{2+} + 2Hg^{2+} \rightarrow Sn^{4+} + Hg^{2+}$

Here Sn^{2+} gets oxidised while Hg^{2+} gets reduced.

9. Zero

10.

$$\overset{x+1}{(NH_4)_2}\,\overset{-2}{SO_4}$$

$$2x + 8(+1) - 2 = 0$$
$$2x + 8 - 2 = 0$$
$$x = -3$$

The oxidation number of N is $(NH_4)_2 SO_4$ is – 3.

Short Answer:

Ans: 1. Ox. No. of N in HNO_3 = +5

Ox No. of N in HNO_2 = +3

Maximum oxidation numbers which N can show is = + 5

($\therefore$ It has only 5 valance electrons $2S^2 2P^3$)

The Ox. No. of N in HNO_3 is maximum and it can only decrease. Therefore, HNO_3 can act only as an oxidant. Minimum Ox. No. of N is -3.

Thus, HNO_2 in which Ox. No. of N is +3 Can decrease as well as increase. Thus, HNO_2 can act as an oxidant as well as a reductant.

Ans: 2. Oxidation half reaction

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

$$Zn(s) \longrightarrow Zn^{2+} + 2e^- \qquad \ldots(1)$$

(To balance charge)

Reduction half reaction

$$NO_3^-(aq) \longrightarrow NH_4^+(aq)$$

$$NO_3^-(aq) \longrightarrow NH_4^+ + 3H_2O(l) \quad \text{(to balance O atom)}$$

$$NO_3^-(aq) + 10H^+(aq) \longrightarrow NH_4^+(aq) + 3H_2O(l)$$

(to balance H atom)

$$NO_3^-(aq) + 10H^+(aq) + 8e^- \longrightarrow NH_4^+ + 3H_2O(l)$$

Multiply (1) equation by 4 to equalize the no. of electron in both. Add both half reaction

4 Zn(s) → Zn^2+(aq) + 8e–

Zn(s) + NO_3– (aq) + 10H+ (aq) → $4Zn^{2+}$ (aq) +NH^+(aq)+ $3H_22O(1)$

– 1 + 10 = + 8 + 1

+ 9 = + 9

Ans: 3.

increase in oxidation number

$$\overset{+7}{MnO_4^-}(aq) + \overset{+3}{COOH} \longrightarrow \overset{+2}{Mn}(aq) + \overset{+4}{CO_2}(g) + H_2O(l)$$

$$\begin{array}{l} | \\ COOH \end{array}$$

decrease in oxidation number

$$2\,MnO_4^-(aq) + 5\begin{array}{l} COOH(aq) \\ | \\ COOH \end{array} \longrightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$$

$$2\,MnO_4^-(aq) + 5\begin{array}{l} COOH(aq) \\ | \\ COOH \end{array} + 6H^+(aq) \longrightarrow$$

$$2Mn^{2+}(aq) + 10CO_2 + 8H_2O$$

Ans: 4. (i) 2Mg + SO_2 → 2MgO + S

Mg = Reducing agent

SO_2 = Oxidising agent

(ii)$2Cu^{2+} + 41 \rightarrow 2CuI + I_2$

Cu^{2+} = Oxidising agent

I- = Reducing agent

(iii)$SO_2 + 2H_2S \rightarrow 2H_2O + 3S$

SO_2 = Oxidising ageiit

H_2S = Reducing agent

(iv)$Sn^{2+} + 2Hg^{2+} \rightarrow Hg2^{2+} + Sn^{4+}$

Sn^{2+} = Reducing agent

Hg^{2+} = Oxidising agent

Ans: 5. (i) $Zn \rightarrow Zn^{2+} + 2e^-$

Oxidation

(ii) $Cl2 + 2e^- \rightarrow 2Cl-$

Reduction

(iii) $Fe \rightarrow Fe^{2+} + 2e^-$

Oxidation

(tv) $Sn^{4+} + 2e^- \rightarrow Sn^{2+}$

Reduction

Ans: 6. The minimum oxidation number shown by S is – 2 since it can acquire 2 more electrons to achieve the nearest inert gas [Ar] configuration.

The maximum Ox. No. shown by S is +6 since it has 6 valance electrons. ($3S^2$ $3P^4$)

Long Answer:

Ans: 1. Various atoms are assigned oxidation number on the basis of the following rules:

1. An element in the free state has an oxidation number equal to zero, e.g. H2, He, K, Ag all have zero ox. no.

2. In a binary compound of a metal and a non-metal, the oxidation number of metal is positive while that of non-metal is negative. In NaCl the ox. no. of sodium +1 and ox. n. of chlorine is -1.

3. In a covalent compound, the atom with higher electronegativity has a negative oxidation number while another atom has a positive oxidation number.

4. The oxidation number of the radical or ions is equal to the electrical charge on it. for e.g. the ox. no. of Na+ is +1.

5. In neutral molecules, the algebraic sum of the oxidation number of all the atoms is zero.

Ans: 2. The skeletal equation is
$MnO_4^- + H^+ + Cl^-(aq) \rightarrow Mn^{2+} + Cl_2(g) + H_2O\ (l)$

Ox. no. of Mn change from +7 in MnO_4^- to +2
Whereas ox. no. of chlorine changes from -1 in Cl^- ions to 0 in Cl_2.

$$Mn \downarrow 5 \times 2$$

$$2Cl \longrightarrow Cl_2 \uparrow 2 \times 5$$

$$2\,MnO_4^-(aq) + 10Cl^-(aq) \longrightarrow 2Mn^{2+}(aq) + 5Cl_2(g)$$

(To balance 0 atoms)

$$2\,MnO_4^-(aq) + 10Cl^-(aq) \longrightarrow 2Mn^{2+}(aq) + 5Cl_2 + 8H_2O$$

$$2\,MnO_4^-(aq) + 10Cl^-(aq) + 16H^+ \longrightarrow 2Mn^{2+} + 5Cl_2 + H_2O$$

(to balance H atoms)

Ans: 3. The following steps are used for balancing the reactions by this methods:

1. Writing the skeletal equation for all the reactants and products of the reaction.
2. Assignment of the oxidation number of all atoms in each compound in the skeletal equation. Identify the atoms undergoing a change in their oxidation number.
3. Calculating the increase or decrease in oxidation number per atom and then for the whole molecule in which it occurs. If these are not equal, then multiplying by suitable coefficients such that these become equal.
4. Now balancing the chemical reaction with respect to all atoms except H & O.

5. Finally balancing with respect to H & O atom for balancing oxygen atoms add H20 molecules to the side deficient in it.

Ans: 4. (i) OF_2

Let the ox. no. of O = x

The ox. no. of each F = — 1

x – 2 = 0

x = +2

(ii) Na_2O_2

Let the o. no. of O = x

ox. no. of each Na = + 1

2 + 2x = 0

2x = – 2

x = – 1

(iii) CH_3COOH

Let the ox. no. of O = x

The ox. no. of each carbon atom = — 1

The ox. no. of hydrogen = +1

– 2 + 4 + 2x = 0

2x + 2 = 0

x = – 1

Ans: 5. The balanced ionic equation for the reaction is

$MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

from the balanced equation, it is evident that-

1 mole of $KMnO_4$ = 5 moles of $FeSO_4$

Applying the molarity equation to the balanced redox equation.

$$\frac{M_1V_1}{n_1}(KMnO_4) = \frac{M_2V_2}{n_2}(FeSO_4)$$

or $$\frac{1 \times V_1}{8 \times 1} = \frac{1}{4} \times \frac{25}{5}$$

or $$V_1 = \frac{1 \times 25 \times 8}{4 \times 5} = 10.0 \text{ cm}^3$$

Thus, the volume of M/8 $KMnO_4$ solution required = 10.0 ml.

Assertion Reason Answer:

1. (ii) Both A and R are true but R is not the correct explanation of A.
2. (iii) A is true but R is false.

Case Study Answer:

1. Answer:

(1) (c) 3

(2) (b) Evolved

(3) (b) Stronger reducing agent than H^+/H_2 couple

(4) (b) In redox reaction, the oxidation number of oxidant increases, while that of reductant decreases

2. Answer:

(1) (a)$KMnO_4$

(2) (c) Fluorine

(3) (c) + 6

(4) (b) + 3