Chapter 8 Redox Reactions

Redox Reaction Basic Formulas

Oxidation	Reduction
1. Addition of oxygen	1. Removal of oxygen
2. Removal of hydrogen	2. Addition of hydrogen
3. Addition of an electronegative element	3. Removal of an electronegative element
4. Removal of an electropositive element	4. Addition of an electropositive element
5. Loss of electron	5. Gain of electron

Oxidation number denotes theoxidation state of an element in a compound ascertained according to a setof rules formulated on the basis thatelectron in a covalent bond belongsentirely to more electronegative element.

Calculation of oxidation number

1. O. S. of all the elements in their elemental form (in standard state) is taken as zero O. S. of elements in Cl_2, F_2, O_2, P_4, O_3, $Fe_{(s)}$, H_2, N_2, C(graphite) is zero.

2. Common O. S. of elements of group one (1^{st}) is one. Common O. S. of elements of group two (2^{nd}) is two.

3. For ions composed of only one atom, theoxidation number is equal to the chargeon the ion.

4. The oxidation number of oxygen in most compounds is –2.While in peroxides (e.g., H_2O_2, Na_2O_2), eachoxygen atom is assigned an oxidationnumber of –1, in superoxides (e.g., KO,RbO_2) each oxygen atom is assigned anoxidation number of –(½).

5. In oxygendifluoride (OF2) and dioxygendifluoride (O_2F_2), the oxygen is assignedan oxidation number of +2 and +1,respectively.

6. The oxidation number of hydrogen is +1 but in metal hydride its oxidation no. is–1.

7. In all its compounds, fluorine has anoxidation number of –1.

8. The algebraic sum of the oxidation numberof all the atoms in a compound must bezero.

9. In polyatomic ion, the algebraic sumof all the oxidation numbers of atoms ofthe ion must equal the charge on the ion.

Stockotation:the oxidationnumber is expressed by putting a Romannumeral representing the oxidation numberin parenthesis after the symbol of the metal inthe molecular formula. Thus aurous chlorideand auric chloride are written as $Au_{(I)}$Cl andAu(III)Cl3. Similarly, stannous chloride andstannic chloride are written as Sn(II)Cl2andSn(IV)Cl_4.

Oxidation: An increase in the oxidationnumber

Reduction: A decrease in the oxidationnumber

Oxidising agent: A reagent which canincrease the oxidation number of an elementin a given substance. These reagents are calledas **oxidants** also.

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

Redox reactions: Reactions which involvechange in oxidation number of the interactingspecies

Balancing of redox reactions:

Oxidation Number Method:

Write the net ionic equation for the reaction of potassium dichromate(VI), $K_2Cr_2O_7$ with sodium sulphite,Na2SO_3, in an acid solution to give chromium(III) ion and the sulphate ion.

Step 1: The skeletal ionic equation is: $Cr_2O_7^{2-}$(aq) + SO_3^{2-}(aq) → Cr_3+(aq)+ SO_4^{2-}(aq)

Step 2: Assign oxidation numbers forCr andS+6 –2 +4 –2 +3 +6 –2 $Cr_2O_7^{2-}$(aq) + SO_3^{2-}(aq) → Cr_3+(aq)+ SO_4^{2-}(aq)

Step 3: Calculate the increase anddecrease of oxidation number, and make them equal:+6 –2 +4 –2 +3 +6 $Cr_2O_7^{2-}$(aq) + 3SO_3^{2-}(aq) → 2Cr_3+(aq)+ 3SO_4^{2-}(aq)

Step 4: Balance the charge by adding H+as the reaction occurs in theacidic medium, $Cr_2O_7^{2-}$(aq) + 3SO_3^{2-}(aq) 8H+ → 2Cr_3+(aq)+ 3SO_4^{2-}(aq)

Step 5: Balance the oxygen atom by adding water molecule. $Cr_2O_7^{2-}$(aq) + 3SO_3^{2-}(aq) 8H+→ 2Cr_3+(aq)+ 3SO_4^{2-}(aq)+ 4H_2O(l)

Half Reaction Method

balance the equation showing the oxidation of Fe_2.+ ions to Feb_3+ ions by dichromate ions ($Cr_2O_7^{2-}$ in acidic medium,wherein, Cr_2 O_7^{2-} ions are reduced to Cr_3+ ions.

Step 1: Produce unbalanced equation for thereaction in ionic form: Fe2+(aq) + $Cr_2O_7^{2-}$ (aq) → Fe_3+ (aq) + Cr_3+(aq)

Step 2: Separate the equation into halfreactions: +2 +3 Oxidation half : Fe2+ (aq) → Fe3+(aq) +6 –2 +3 Reduction half :$Cr_2O_7^{2-}$(aq) → Cr_3+(aq)

Step 3: Balance the atoms other than O andH in each half reaction individually. $Cr_2O_7^{2-}$ (aq) → Cr_3+(aq)

Step 4: For reactions occurring in acidicmedium, add H2O to balance O atoms and H+to balance H atoms.$Cr_2O_7^{2-}$ (aq) +14 H+→ Cr_3+(aq) + 7H_2O (l)

Step 5: Add electrons to one side of the halfreaction to balance the charges. If need be,make the number of electrons equal in the twohalf reactions by multiplying one or both halfreactions by appropriate coefficients. Fe_2+ (aq) → Fe_3+ (aq) + e– $Cr_2O_7^{2-}$ (aq) + 14H+ (aq) + 6e– → 2Cr_3+(aq) +7H_2O (l) 6Fe_2+ (aq) →6 Fe_3+ (aq) +6 e–

Step 6: We add the two half reactions toachieve the overall reaction and cancel theelectrons on each side. This gives the net ionicequation as : 6Fe_2+(aq) + $Cr_2O_7^{2-}$ (aq) + 14H+(aq) → 6 Fe_3+(aq) +2Cr 3+(aq) + 7H_2O(l)

A **redox couple** is defined as havingtogether the oxidised and reduced forms of asubstance taking part in an oxidation orreduction half reaction. Represented as Zn_2+/Zn and Cu_2+/Cu.

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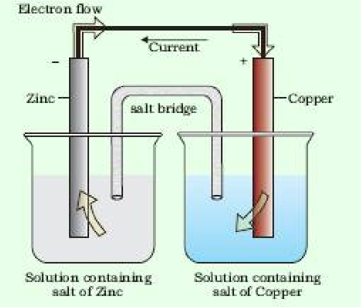

The potential associated with eachelectrode is known as **electrode potential**. Ifthe concentration of each species taking partin the electrode reaction is unity (if any gasappears in the electrode reaction, it is confinedto 1 atmospheric pressure) and further thereaction is carried out at 298K, then thepotential of each electrode is said to be the **Standard Electrode Potential**

- SHE is used to measure electrode potential and its standard electrode potential is taken as 0.00 V.

CHAPTER-8
REDOX REACTIONS

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5. In oxygendifluoride (OF_2) and dioxygendifluoride (O_2F_2), the oxygen is assignedan oxidation number of +2 and +1,respectively.
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Step 2: Assign oxidation numbers forCr and S

+6 –2 +4 –2 +3 +6 –2

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

Step 3: Calculate the increase anddecrease of oxidation number, and make them equal:

+6 –2 +4 –2 +3 +6

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

Step 4: Balance the charge by adding H^+as the reaction occurs in theacidic medium,

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

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balance the equation showing the oxidation of Fe^{2+} ions to Fe^{3+} ions by dichromate ions $(Cr_2O_7)^{2-}$ in acidic medium,wherein, $Cr_2O_7^{2-}$ ions are reduced to Cr^{3+} ions.

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Step 2: Separate the equation into halfreactions:

+2 +3

Oxidation half : $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq)$

+6 –2 +3

Reduction half :$Cr_2O_7^{2-}(aq) \rightarrow Cr^{3+}(aq)$

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Step 5: Add electrons to one side of the halfreaction to balance the charges. If need be,make the number of electrons equal in the twohalf reactions by multiplying one or both halfreactions by appropriate coefficients.

Fe^{2+} (aq) → Fe^{3+} (aq) + e–

$Cr_2O_7^{2-}$ (aq) + $14H^+$ (aq) + 6e– → $2Cr^{3+(}$aq) +$7H_2O$ (l)

$6Fe^{2+}$ (aq) →6 Fe^{3+} (aq) +6 e–

Step 6: We add the two half reactions toachieve the overall reaction and cancel theelectrons on each side. This gives the net ionicequation as :

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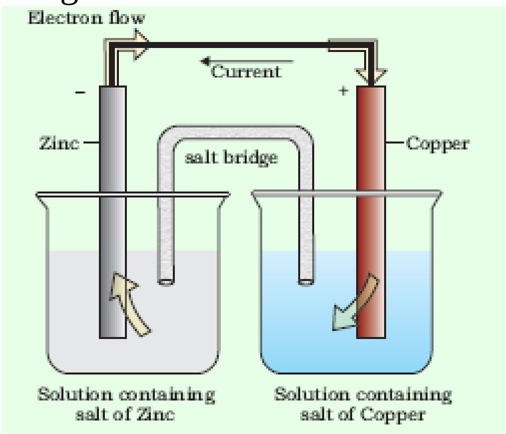
A **redox couple** is defined as havingtogether the oxidised and reduced forms of asubstance taking part in an oxidation orreduction half reaction.

Represented as $Zn^{2+/}$Zn and Cu^{2+}/Cu.

❖ Electrochemical cells are the devices which are used to get electric current by using chemical reaction.



Daniell cell having electrodes of zinc and copper dipping in the solutions of their respective salts.

The potential associated with eachelectrode is known as **electrode potential**. Ifthe concentration of each species taking partin the electrode reaction is unity (if any gasappears in the electrode reaction, it is confinedto 1 atmospheric pressure) and further thereaction is carried out at 298K, then thepotential of each electrode is said to be the**Standard Electrode Potential**.

- SHE is used to measure electrode potential and its standard electrode potential is taken as 0.00 V.

<u>ONE MARK QUESTIONS</u>

1. Define oxidation and reduction in terms of oxidation number.
   Ans Increase in Oxidation Number is Oxidation and decrease in Oxidation Number is called reduction.
2. What is meant by disproportionation? Give one example.
   Ans : In a disproportionation reaction an element simultaneously oxidized and reduced.
   $P_4 + 3OH^- + 3H_2O \rightarrow PH_3 + 3H_2PO_2^-$
3. What is O.N. of sulphur in H_2SO_4?Ans: +6

4. Identify the central atom in the following and predict their O.S.
   HNO_3
   Ans: central atom:- N; O.S. +5
5. Out of Zn and Cu which is more reactive?
   Ans: Zn.
6. What is galvanization?
   Ans: Coating of a less reactive metal with a more reactive metal e.g.- coating of iron surface with Zn to prevent rusting of iron.
7. How is standard cell potential calculated using standard electrode potential?
   Ans: $E^0_{cell} = E^0_{cathode} - E^0_{anode}$
8. What is O.S. of oxygen in H_2O_2?
   Ans: - -1.
9. The formation of sodium chloride from gaseous sodium and gaseous chloride is a redox process justify.
   Ans: Na atom get oxidize and Cl is reduced.

TWO MARKS QUESTIONS

1. Write the balanced redox reaction .
   (I) $MnO_4^-(aq) + Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + Fe^{3+}(aq)$ [acidic medium]
   (II) $Cr_2O_7^{2-} + Fe^{2+} \rightarrow Cr^{3+} + Fe^{3+}$ [Acidic medium]

   Ans:- (i) $MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O_{(l)}$
   (ii) $Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$
2. Identify the strongest & weakest reducing agent from the following metals: .Zn, Cu, Na, Ag, Sn
   Ans: Strongest reducing agent: Na, weakest reducing agent: Ag.
3. Determine the oxidation no. of all the atoms in the following oxidants:$KMnO_4$, $K_2Cr_2O_7$ and $KClO_4$
   Ans :
   In $KMnO_4$ K = +1, Mn = +7, O = -2
   In $K_2Cr_2O_7$K = +1, Cr = +6, O = -2
   In $KClO_4$K = +1, Cl = =+7, O = -2
4. Determine the oxidation no. of all the atoms in the following species:Na_2O_2 and OF_2.
   Ans: In Na_2O_2Na = +1, O = -1
   InOF_2, F = -1, O = +2

5. Is it possible to store :
   (i) H_2SO_4 in Al container?(ii) CuSO4 solution in Zn vessel?
   Ans : (i) yes. (ii) No.
6. Calculate the standard e.m.f. of the cell formed by the combination of $Zn/Zn^{2+} || Cu^{2+}/Cu$.
   Solution- : $E^0_{cell} = E^0_{cathode} - E^0_{anode}$
   $=0.34 - (-0.76) = 1.10V.$
7. Identify the oxidizing and reducing agents in the following equations:

(i) $MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O_{(l)}$

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

Ans : (i) O.A. = MnO_4^- ; R.A.= Fe^{2+}

(ii)O.A.=$Cr_2O_7^{2-}$; R.A.= Fe^{2+}

8. Predict all the possible oxidation states of Cl in its compounds.
   Ans:- 0, -1, +1, +3, +5, +7
9. Formulate possible compounds of 'Cl' in its O.S.is: 0, -1, +1, +3, +5, +7
   Ans: Cl_2, HCl, HOCl, HOClO, $HOClO_2$, $HOClO_3$ respectively.
10. List three measures used to prevent rusting of iron.
   Ans: (i) galvanization(coating iron by a more reactive metal)
   (ii) greasing/oiling
   (iii) painting.

THREE MARK QUESTIONS

1. Write short notes on :
   (a) Electrochemical series(b) redox reactions (c) oxidizing agents
   Ans :(a) Electrochemical series :- arrangement of metals(non-metals also) in increasing order of their reducing power or vice versa.
   (b) Reactions in which both Oxidation and reduction take place simultaneously are REDOX REACTIONS.
   (c)oxidizing agents : chemical specie which can oxidize the other one or can reduce itself.
2. Calculate O. S. of sulphur in the following oxoacids of 'S' :
   H_2SO_4 ,$H_2SO_3H_2S_2O_8$and $H_2S_2O_7$
   Ans : +6, +4, +6 and +6 respectively.
   (calculate by considering x of 'S' and taking +1 of H, -2 0f "O" and -1 of "O" in peroxide bond.)
3. Explain role of salt bridge in Daniell cell.
   Ans : (a) it completes the electric circuit in the cell.
   (b) it maintains the electric neutrality in the cell.

4. Account for the followings :
   (i) sulphur exhibits variable oxidation states.
   Ans. Due to the presence of vacant 'd' orbitals in 'S'
   (ii) Fluorine exhibits only -1 O.S.
   Ans . It is most electronegative element
   (iii) oxygen can't extend its valency from 2.
   Ans. Small size/unavailabilityof vacant 'd' orbitals in O
5. Balance the equation $MnO_4^- + I^- \rightarrow Mn^{2+} + I_2 + H_2O$by ion electron method in acidic medium.
   Ans :<u>Step-I</u> Balancing of reduction half reaction by adding protons and electrons on LHS and more water molecules on RHS:
   $8H^+ + MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$

Step-II Balancing of oxidation half reaction by adding electrons on RHS:
$2I^- \rightarrow I_2 + 2e^-$
Step-III To multiply the OHR by 5; RHR by2 andto add OH & RH reactions to get overall redox reaction(cancellation of electrons of RH & OH reactions):

$$[8H^+(aq) + MnO_4^-(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)] \times 2$$

$$[2I^- \rightarrow I_2 + 2e^-] \quad \times \quad 5$$

$$MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O_{(l)}$$

6. complete and balance the following equations:

(i) $H^+ + Cr_2O_7^{2-} + Br^- \rightarrow 2Cr^{3+} + Br_2 +$ -----

(ii) $H_2O_2 + Cl^- \rightarrow OH^- + Cl_2$

(iii) $Zn + Cu^{2+} \rightarrow$?

Ans :(i) $14H^+ + Cr_2O_7^{2-} + 6\,Br^- \rightarrow 2Cr^{3+} + 3Br_2 + 7H_2O$

(ii) $H_2O_2 + 2Cl^- \rightarrow 2OH^- + Cl_2$

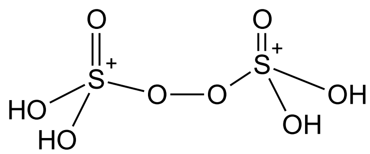
(ii) $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

7. Identify the oxidizing and reducing agents in the following equations:
   (i) $Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$
   (ii) $H_2 + Cl_2 \rightarrow 2HCl$
   (iii) $MnO_2 + 4HCl \rightarrow MnCl_2 + 2H_2O + Cl_2$
   Ans :(i) O.A. $= H_2SO_4$; R.A.= Fe
   (ii) O.A. $= Cl_2$; R.A.$= H_2$
   (iii)O.A. $= MnO_2$; R.A. =HCl
8. . Arrange the following in increasing order of their reducing power:

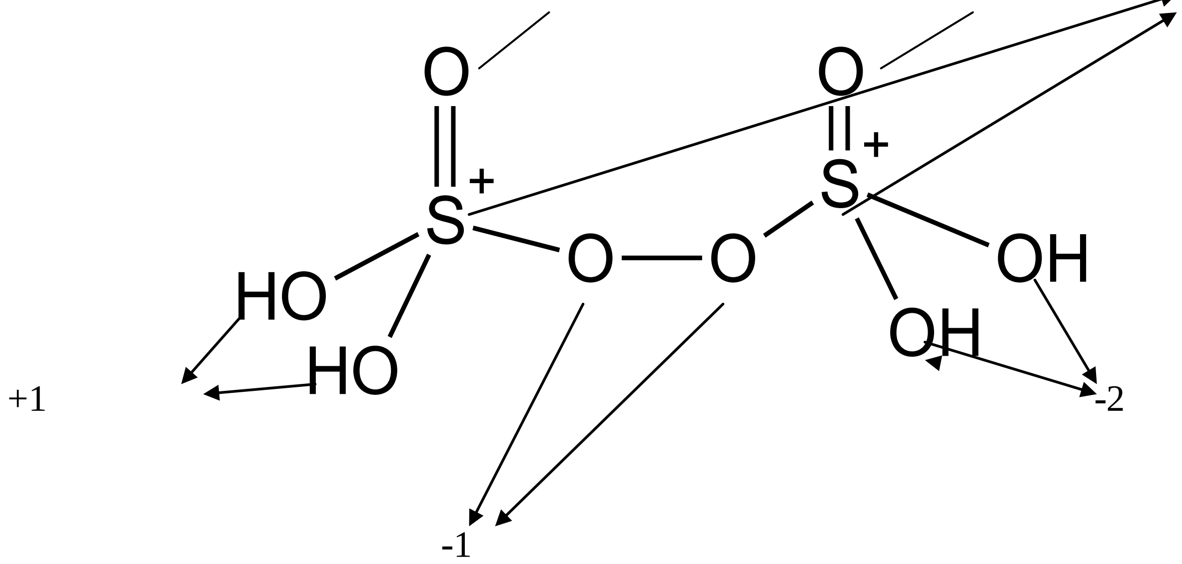
Cu, Ag, Au, Zn, Fe, Al, Na, Mg, Pt(SHE), Hg, Ca, K

Ans : Au, Hg, Ag, Cu, Pt(SHE), Fe, Zn, Al, Mg, Na, Ca, K

9. Indicate O.S. of each atom present in given structure of peroxodisulphuric acid



Ans:



10.What is SHE? What is its use?

Ans :Standard Hydrogen Electrode (SHE) has beenselected to have zero standard potential at alltemperatures. It consists of a platinum foilcoated with platinum black (finely dividedplatinum) dipping partially into an aqueous solution in which the activity (approximateconcentration 1M) of hydrogen ion is unity andhydrogen gas is bubbled through the solutionat 1 bar pressure.The potential of the other half cell is measuredby constructing a cell in which reference electrode is standard hydrogen electrode. The potential of the other half cell is equal to thepotential of the cell.

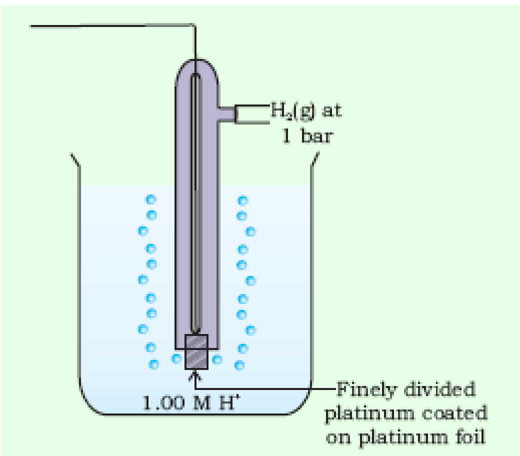


Fig: SHE

HOTS QUESTIONS

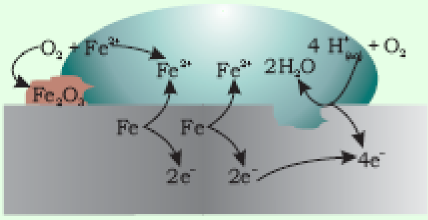
1. Is rusting of iron an electrochemical phenomenon? How ?explain.

Ans : Yes. Rusting of iron is an electrochemical phenomenon because this is possible due to formation of a small electrochemical cell over rough surface of iron and the following redox reaction takes place there in that cell-

Oxidation $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Reduction $O_2 + 4H^+ \ \ 4e^- \ \rightarrow 2H_2O$

$e- + ½ \ O_2 + 2H_2O + 2Fe^{2+} \rightarrow Fe_2O_3 + 4H^+$



Oxidation: $Fe \ (s) \rightarrow \ Fe^{2+} \ (aq) + 2e^-$

Reduction: $O_2 \ (g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$

Atomospheric

oxidation : $2Fe^{2+}(aq) + 2H_2 \ O(l) + ½O_2(g) \rightarrow Fe_2O_3(s) + 4H^+(aq)$

Fig. 5.14 *Corrosion of iron in atmosphere.*

2. We expand croreof Rupees and even thousands of lives every year due to corrosion. How can be preventing it. Explain.

Ans : (i) By Galvanization: Coating of a less reactive metal with a more reactive metal e.g.- coating of iron surface with Zn to prevent rusting of iron.

(ii) By greasing /oiling (to keep away the object from the contact of air & moisture.)

(iii)By painting (to keep away the object from the contact of air & moisture.)

Subject: Chemistry

Class: XI

Chapter: Redox Reactions

Top concepts

1. Redox reactions are those reactions in which oxidation and reduction takes place simultaneously

2. Classical view of redox reactions
   - Oxidation is addition of oxygen / electronegative element to a substance or removal of hydrogen / electropositive element from a substance
   - Reduction is removal of oxygen / electronegative element from a substance or addition of hydrogen / electropositive element to a substance

3. Redox reactions in terms of Electron transfer
   - Oxidation is defined as loss of electrons by any species
   - Reduction is defined as gain of electrons by any species

4. In oxidation reactions there is loss of electrons or increase in positive charge or decrease in negative charge

5. In reduction reactions there is gain of electrons or decrease in positive charge or increase in negative charge

6. Oxidising agents are species which gain one or more electrons and get reduced themselves

7. Reducing agents are the species which lose one or more electrons and gets oxidized themselves

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

9. Rules for assigning oxidation number to an atom

- Oxidation number of Hydrogen is always +1 (except in hydrides, it is -1).

- Oxidation number of oxygen in most of compounds is -2. In peroxides it is (-1). In superoxides, it is (-1/2). In OF_2 oxidation number of oxygen is +2.In O_2F_2 oxidation number of oxygen is +1
- Oxidation number of Fluorine is -1 in all its compounds
- For neutral molecules sum of oxidation number of all atoms is equal to zero
- In the free or elementary state, the oxidation number of an atom is always zero. This is irrespective of its allotropic form

- For ions composed of only one atom, the oxidation number is equal t the charge on the ion

- The algebraic sum of the oxidation number of all the atoms in a compound must be zero

- For ions the sum of oxidation number is equal to the charge on the ion

- In a polyatomic ion, the algebraic sum of all the oxidation numbers of atoms of the ion must be equal to the charge on the ion

10. Oxidation state and oxidation number are often used interchangeably

11. According to Stock notation the oxidation number is expressed by putting a Roman numeral representing the oxidation number in parenthesis after the symbol of the metal in the molecular formula

12.Types of Redox Reactions

- Combination Reactions: Chemical reactions in which two or more substances (elements or compounds) combine to form a single substance

- Decomposition Reactions: Chemical reactions in which a compound break up into two or more simple substances

- Displacement Reactions: Reaction in which one ion(or atom)in a compound is replaced by an ion(or atom) of other element

  a. Metal Displacement Reactions: Reactions in which a metal in a compound is displaced by another metal in the uncombined state

  b. Non-metal Displacement Reactions: Such reactions are mainly hydrogen displacement or oxygen displacement reactions

- Disproportionation Reactions: Reactions in which an element in one oxidation state is simultaneously oxidized and reduced

13.Steps involved in balancing a Redox reaction by oxidation number method

- Write the skeletal redox reaction for all reactants and products of the reaction
- Indicate the oxidation number of all the atoms in each compound above the symbol of element
- Identify the element/elements which undergo change in oxidation numbers
- Calculate the increase or decrease in oxidation number per atom
- Equate the increase in oxidation number with decrease in oxidation number on the reactant side by multiplying formula of oxidizing agent and reducing agents with suitable coefficients

- Balance the equation with respect to all other atoms except hydrogen and oxygen
- Finally balance hydrogen and oxygen. For balancing oxygen atoms add water molecules to the side deficient in it. Balancing of hydrogen atoms depend upon the medium
    a. For reactions taking place in acidic solutions add H^+ ions to the side deficient in hydrogen atoms
    b. For reactions taking place in basic solutions add H_2O molecules to the side deficient in hydrogen atoms and simultaneously add equal number of OH^- ions on the other side of the equation
- Finally balance the equation by cancelling common species present on both sides of the equation

14. Steps involved in balancing a Redox by Ion-Electron Method(Half reaction method)

- Find the elements whose oxidation numbers are changed. Identify the substance that acts as an oxidizing agent and reducing agent
- Separate the complete equation into oxidation half reaction and reduction half reaction
- Balance the half equations by following steps
    i. Balance all atoms other than H and O
    ii. Calculate the oxidation number on both sides of equation .Add electrons to whichever side is necessary to make up the difference
    iii. Balance the half equation so that both sides get the same charge
    iv. Add water molecules to complete the balancing of the equation
- Add the two balanced equations. Multiply one or both half equations by suitable numbers so that on adding two equations the electrons are balanced

15. Application of Redox reactions: Redox Titrations

- Potassium permanganate in redox reactions: Potassium permanganate ($KMnO_4$) is very strong oxidizing agent and is used in determination of many reducing agents like Fe^{2+}, oxalate ions etc. It acts as self indicator in redox reactions.

  Equation showing $KMnO_4$ as an oxidising agent in acidic medium is:

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

- Acidified Potassium dichromate ($K_2Cr_2O_7$) in redox reactions: $K_2Cr_2O_7$ is used as an oxidizing agent in redox reactions. Titrations involving $K_2Cr_2O_7$ uses diphenylamine and potassium ferricyanide (external indicator).

  Equation showing $K_2Cr_2O_7$ as an oxidising agent in acidic medium is:

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

- Iodine (I_2) in redox reactions: I_2 acts as mild oxidising agent in solution according to equation

  $$I_2 + 2e^- \rightarrow 2I^-$$

16. Direct redox reaction: Redox reactions in which reduction and oxidation occurs in same solution (i.e. same reaction vessel).In these reactions transference of electrons is limited to very small distance.

17. Indirect redox reactions: Redox reactions in which oxidation and reduction reactions take place in different reactions vessels and thus transfer of electrons from one species to another does not take place directly

18. Electrochemical cell is a device that converts chemical energy produced in a redox reaction into electrical energy. These cells are also called Galvanic cells or Voltaic cells

19. The electrode at which oxidation occurs is called anode and is negatively charged

20. The electrode at which reduction takes place is called cathode and is positively charged

21. In an electrochemical cell the transfer of electrons takes place from anode to cathode

22. In an electrochemical cell the flow of current is from cathode to anode

23. In the electrochemical cell, the electrical circuit is completed with a salt bridge. Salt bridge also maintains the electrical neutrality of the two half cells

24. A salt bridge is a U shaped tube filled with solution of inert electrolyte like sodium chloride or sodium sulphate which will not interfere in the redox reaction. The ions are set in a gel or agar agar so that only ions flow when inverted

25. Electrical potential difference developed between the metal and its solution is called electrode potential. It can also be defined as tendency of an electrode in a half cell to gain or lose electrons

26. Oxidation potential is the tendency of an electrode to lose electrons or to get oxidized

27. Reduction potential is the tendency of an electrode to gain electrons or get reduced

28. In an electrochemical cell, by the present convention, the electrode potentials are represented as reduction potential

29. The electrode having a higher reduction potential will have a higher tendency to gain electrons

30. By convention, the standard electrode potential of hydrogen electrode is 0.00 volts

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

32. The difference between the electrode potentials of eth two electrodes constituting the electrochemical cell is called EMF(Electromotive force) or the cell potential

33. EMF = $E^{\Theta}_{cathode} - E^{\Theta}_{anode}$

34. A negative E^{θ} means that the redox couple is a stronger reducing agent than the H^+/H_2 couple

35. A positive E^{θ} means that the redox couple is a weaker reducing agent than the H^+/H_2 couple