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Class 11 Redox Reactions Important Questions with Answers

Short Answer Type Questions

Q1. The reaction

$Cl_2(g) + 2OH^-(aq) \rightarrow ClO^-(aq) + Cl^-(aq) + H_2O(l)$

represents the process of bleaching. Identify and name the species that bleaches the substances due to their oxidising action.

Answer:

$$Cl_2(g) + 2OH^-(aq) \rightarrow ClO^-(aq) + Cl^-(aq) + H_2O(l)$$

Here, Cl_2 is both oxidized and reduced in ClO^- and Cl^-, respectively. Since Cl^- cannot act as an oxidising agent (O.A.). Therefore, Cl_2 bleaches substances due to oxidising action of hypochlorite ClO^- ion.

Q2. MnO_4^{2-} undergoes a disproportionation reaction in an acidic medium but MnO_4^- does not. Give a reason.

Answer:

Disproportionation is a redox reaction in which one intermediate oxidation state component transforms into two higher and lower oxidation state compounds.

Manganese's oxidation states ranging from +2 to +7 in its various compounds. MnO_4^- has the maximum oxidation state of +7 hence disproportionation is impossible, but MnO_4^{2-} has a +6 oxidation state, which can be oxidised as well as reduced.

Q3. PbO and PbO_2 react with HCl according to following chemical equations:

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

Why do these compounds differ in their reactivity?

Answer:

None of the atoms' O.N. changes in the first reaction. As a result, it cannot be classified as a redox reaction. Because PbO is a basic oxide that combines with HCl acid, it is an acid-base reaction.

PbO_2 is reduced and functions as an oxidising agent in the second process, a redox reaction.

Q4. Nitric acid is an oxidising agent and reacts with PbO, but it does not react with PbO_2. Explain why?

Answer:

PbO is a basic oxide, and it undergoes a straightforward acid-base reaction with HNO_3. In PbO_2, on the other hand, lead is in the +4 oxidation state and cannot be further oxidised. As a result, no reaction occurs. PbO_2 is thus inactive; only PbO interacts with HNO_3.

$$2PbO + 4HNO_3 \rightarrow 2Pb(NO_3)_2 + 2H_2O$$

Q5. Write a balanced chemical equation for the following reactions:

(i) Permanganate ion (MnO_4^-) reacts with sulphur dioxide gas in an acidic medium to produce Mn^{2+} and hydrogensulphate ion.
(Balance by ion electron method)

(ii) Reaction of liquid hydrazine (N_2H_4) with chlorate ion (ClO_3^-) in basic medium produces nitric oxide gas and chloride ion in a gaseous state.
(Balance by oxidation number method)

(iii) Dichlorine heptaoxide (Cl_2O_7) in gaseous state combines with an aqueous solution of hydrogen peroxide in acidic medium to give chlorite ion (ClO_2^-) and oxygen gas.
(Balance by ion electron method)

Answer:

(i) $MnO_4^- + SO_2 \rightarrow Mn^{2+} + HSO_4^-$ (acidic medium)

Balancing by ion-electron method we get:

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

5 * {$SO_2 + 2H_2O \rightarrow HSO_4^- + 3H^+ + 2e^-$

$$
\begin{array}{l}
2MnO_4^- + 16H^+ + \cancel{10e^-} \longrightarrow 2Mn^{2+} + 8H_2O \\
\qquad 5SO_2 + 10H_2O \longrightarrow 5HSO_4^- + 15H^+ + \cancel{10e^-} \\
\hline
2MnO_4^- + H^+ + 5SO_2 + 2H_2O \longrightarrow 2Mn^{2+} + 5HSO_4^-
\end{array}
$$

(ii) $N_2H^4 + ClO_3^- \rightarrow NO + Cl^-$ (basic medium)

To make the gain and loss of electrons equal, balancing by oxidation number method we get:

$3N_2H_4 + 4ClO_3^- \rightarrow 6NO + 4Cl^- + 6H_2O$

decrease by 6×4

$$\overset{-2}{N_2}\ \overset{+1}{H_4} + \overset{+5}{Cl}O_3^- \longrightarrow 2\overset{+2}{N}O + Cl^-$$

increase by 8×3

(iii) $Cl_2O_7(g) + H_2O_2(aq) \rightarrow 2ClO_2^- + 3H_2O$ (acidic medium)

2 * {$Cl_2O_7 + 6H^+ + 8e^- \rightarrow 2ClO_2^- + 3H_2O$

8 * {$H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$

$$
\begin{array}{l}
2Cl_2O_7 + 12H^+ + \cancel{16e^-} \longrightarrow 4ClO_2^- + 6H_2O \\
\qquad\qquad 8H_2O_2 \longrightarrow 8O_2 + 12H^+ + \cancel{16e^-} \\
\hline
\qquad 2Cl_2O_7 + 8H_2O_2 \longrightarrow 4ClO_2^- + 6H_2O + 8O_2 + 4H^+
\end{array}
$$

Q6. Calculate the oxidation number of phosphorus in the following species.
(a) HPO_3^{2-} and
(b) PO_4^{3-}

Answer:

(a) Let the oxidation number of phosphorus in HPO_3^{2-} be x.

$H + P + 3O^{2-}$
⇒ +1 + x + (-2)*3 = -2
⇒ +1 + x - 6 = -2
⇒ x - 5 = -2
⇒ x = - 2 + 5
⇒ x = +3

Thus, the oxidation number of phosphorus is **+3**.

(b) Let the oxidation number of phosphorus in PO_4^{3-} be x.

PO_4^{3-}
⇒ x + 4 * (-2) = -3
⇒ x = -3 + 8 = +5
⇒ x = +5

Thus, the oxidation number of phosphorus is **+5**.

Q7. Calculate the oxidation number of each sulphur atom in the following compounds:

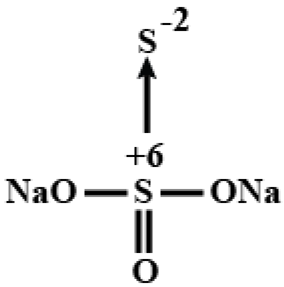
(a) $Na_2S_2O_3$
(b) $Na_2S_4O_6$
(c) Na_2SO_3
(d) Na_2SO_4

Answer:

(a) The inorganic chemical sodium thiosulfate (also known as sodium thiosulphate) has the formula $Na_2S_2O_3$ xH_2O.

We can't find the oxidation number of each sulphur atom using traditional methods.

Structure will provide a clear explanation. One sulphur molecule possesses +6 O.S., whereas the other has -2.



(b) Let the oxidation number of sulphur in $Na_2S_4O_6$ be x.

$\Rightarrow 2 + 4x - 12 = 0$
$\Rightarrow 4x = +10$
$\Rightarrow x = 10/4$
$\Rightarrow x = +2.5$

Thus, the oxidation number of sulphur in $Na_2S_4O_6$ is +2.5.

(c) Let the oxidation number in Na_2SO_3 be x.

$Na_2\ S\ O_3$

$\Rightarrow 2\times(+1)+x+(-2)\times3 = 0$
$\Rightarrow 2+x-6 = 0$
$\Rightarrow x = +4$

Thus, the oxidation number of sulphur in Na_2SO_3 is +4.

(d) Let the oxidation number of sulphur in Na_2SO_4 be x.

$\Rightarrow 2\times(+1)+x+(-2)\times4 = 0$
$\Rightarrow 2+x-8 = 0$
$\Rightarrow x = +6$

Thus, the oxidation number of sulphur in Na_2SO_4 is +6.

Q8. Balance the following equations by the oxidation number method.

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

(ii) $I_2 + NO_3^- \longrightarrow NO_2 + IO_3^-$

(iii) $I_2 + S_2O_3^{2-} \longrightarrow I^- + S_4O_6^{2-}$

(iv) $MnO_2 + C_2O_4^{2-} \longrightarrow Mn^{2+} + CO_2$

Answer:

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

The chromium oxidation number drops from +6 in $Cr_2O_7^2$ to +3 in Cr^{3+}. In $Cr_2O_7^{2-}$ to Cr^{3+}, the total drop for two chromium atoms is 6. Iron's oxidation number, on the other hand, rises from +2 (in Fe^{2+}) to +3 (in Fe^{3+}).

Multiply Fe2+ by 6 and Cr2O72- by 1 to balance the increase and reduction in oxidation numbers. Then there's,
$6Fe^{2+} + Cr_2O_7^{2-} + H^+ \rightarrow Fe^{3+} + Cr^{3+} + H_2O$

On both sides of the reaction, Fe and Cr atoms must be balanced.
$6Fe^{2+} + Cr_2O_7^{2-} + H^+ \rightarrow 6Fe^{3+} + 2Cr^{3+} + H_2O$

Multiply H_2O by 7 to balance O atoms.
$6Fe^{2+} + Cr_2O_7^{2-} + H^+ \rightarrow 6Fe^3+ + 2Cr^{3+} + 7H_2O$

Multiply H+ by 14 to balance H-atoms.
$6Fe^{2+} + Cr_2O_7^{2-} + 14H^+ \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$

(ii) The given skelton equation is: $I_2 + NO_3^- \rightarrow NO_2 + IO_3^-$

Let us find the oxidation number of both I and N.

LHS,
I_2
$\Rightarrow 2x = 0$
$\Rightarrow x = 0$

IO_3^-
$\Rightarrow x + 3(-2) = -1$
$\Rightarrow x = +5$

For one atom of I, the oxidation number rises by 5. It's 10 for two atoms.

NO_3^-
$\Rightarrow x + 3(-2) = -1$
$\Rightarrow x = +5$

NO_2
$\Rightarrow x + 2(-2) = 0$
$\Rightarrow x = +4$

As a result, the oxidation number drops by 1.

Cross multiplying the oxidation number and writing the equation,
$I_2 + 10NO_3^- \rightarrow NO_2 + IO^{3-}$

Balancing the RHS side I and N,
$I_2 + 10NO_3^- \rightarrow 10NO_2 + 2\ IO_3^-$

Now, balancing the the oxygen on both the sides by adding H_2O,
$I_2 + 10NO_3^- \rightarrow 10\ NO_2 + 2IO_3^- + 4H_2O$

Balance the hydrogen by adding H^+,
$I_2 + 10NO_3^- + 8H^+ \rightarrow 10\ NO_2 + 2IO_3^- + 4H_2O$

Hence, the balanced equation is,
$I_2 + 10NO_3^- + 8H^+ \rightarrow 10NO_2 + 2IO_3^- + 4H_2O$

(iii) The unbalanced redox reaction is $I_2 + S_2O_3^{2-} \rightarrow I^- + S_4O_6^{2-}$

Balancing all atoms other than H and O, we get

$I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$

I changes its oxidation number from 0 to -1. The oxidation number of one I atom changes by one. For 2 I atoms, the total change in oxidation number is 2.
The oxidation number of S increases from 2 to 2.5, representing a 0.5 increase in the oxidation number of one S atom.
For four S atoms, the total change in oxidation number is 2.
Increases in oxidation number are counterbalanced by decreases in oxidation number.
The atoms of O are in a state of equilibrium. This is the chemical equation that is balanced.

(iv) $MnO_2 + C_2O_4^{2-} \rightarrow Mn^{2+} + CO_2$

Balancing all atoms, we get:

$$\overset{IV}{Mn}O_2 + \overset{III}{C_2}O_4^{2-} + 4H^+ \rightarrow 2\overset{IV}{C}O_2 + \overset{II}{Mn}^{2+} + 2H_2O$$

The following is an example of an oxidation-reduction (redox) reaction:

$Mn^{IV} + 2e^- \rightarrow Mn^{II}$ (reduction)

$2C^{III} - 2e^- \rightarrow 2C^{IV}$ (oxidation)

$C_2O_4^{2-}$ is a reducing agent, while MnO_2 is an oxidising agent.

Q9. Identify the redox reactions out of the following reactions and identify the oxidising and reducing agents in them.

(i) $3HCl(aq) + HNO_3\,(aq) \longrightarrow Cl_2\,(g) + NOCl\,(g) + 2H_2O\,(l)$

(ii) $HgCl_2\,(aq) + 2KI\,(aq) \longrightarrow HgI_2\,(s) + 2KCl\,(aq)$

(iii) $Fe_2O_3\,(s) + 3CO\,(g) \xrightarrow{\Delta} 2Fe\,(s) + 3CO_2\,(g)$

(iv) $PCl_3\,(l) + 3H_2O\,(l) \longrightarrow 3HCl\,(aq) + H_3PO_3\,(aq)$

(v) $4NH_3 + 3O_2\,(g) \longrightarrow 2N_2\,(g) + 6H_2O\,(g)$

Answer:

(i) Writing the O.N. of on each atom, we get:

$3HCl + HNO_3 \rightarrow Cl_2 + NOCl + 2H_2O$

Cl's O.N. goes from -1 (in HCl) to 0 in this case (in Cl_2). As a result of the oxidation of Cl^-, HCl is used as a reducing agent.

HNO_3 functions as an oxidising agent because the O.N. of N reduces from +5 (in HNO_3) to +3 (in NOCl).

Thus, reaction (i) is, therefore, a redox reaction.

(ii) $HgCl_2 + 2KI \rightarrow HgI + 2KCl$

There is no change in O.N. of any of the atoms, hence this is not a redox reaction.

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

Fe_2O_3 works as an oxidising agent because the O.N. of Fe falls from +3 (in Fe_2O_3) to 0 (in Fe).

CO works as a reducing agent because the O.N. of C increases from +2 (in CO) to +4 (in CO_2). As a result, we have a redox reaction.

(iv) $PCl_3 + 3H_2O \rightarrow 3HCl + H_2PO_3$

There is no change in O.N. of any of the atoms, hence it is not a redox reaction.

(v) $4NH_3 + 3O_2 \rightarrow 2N_2 + 6H_2O$

Because the O.N. of N in N_2 grows from -3 to 0, NH_3 functions as a reducing agent.
Furthermore, because the O.N. of O drops from 0 to O_2 to -2 in H_2O, O_2 functions as an oxidizing agent.
As a result, this is a redox reaction.

Q10. Balance the following ionic equations

(i) $Cr_2O_7^{2-} + H^+ + I^- \longrightarrow Cr^{3+} + I_2 + H_2O$

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

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

(iv) $MnO_4^- + H^+ + Br^- \longrightarrow Mn^{2+} + Br_2 + H_2O$

Answer:

Balanced equations of given unbalanced equations:

(i) $Cr_2O_7^{-2} + 6I^- + 14H^+ \rightleftharpoons 2Cr^{+3} + 3I_2 + 7H_2O$

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

(iii) $2MnO_4^- + 5SO_3^{-2} + 6H^+ \rightleftharpoons 2Mn^{+2} + 5SO_4^{-2} + 3H_2O$

(iv) $2MnO_4^- + 5SO_3^{-2} + 6H^+ \rightleftharpoons 2Mn^{+2} + 5SO_4^{-2} + 3H_2O$

Long Answer Type Questions

Q1. Explain redox reactions on the basis of electron transfer. Give suitable examples.

Answer:

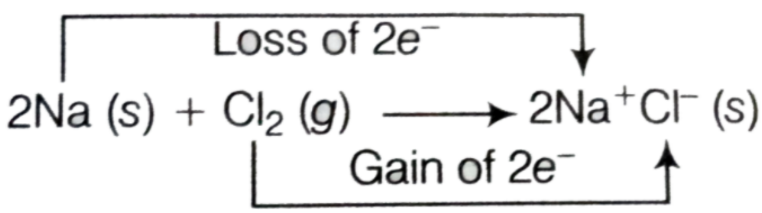
An oxidation-reduction (redox) reaction is a chemical reaction in which electrons are transferred between two substances. Any chemical reaction in which a molecule, atom, or ion's oxidation number changes by acquiring or losing an electron is known as an oxidation-reduction reaction.

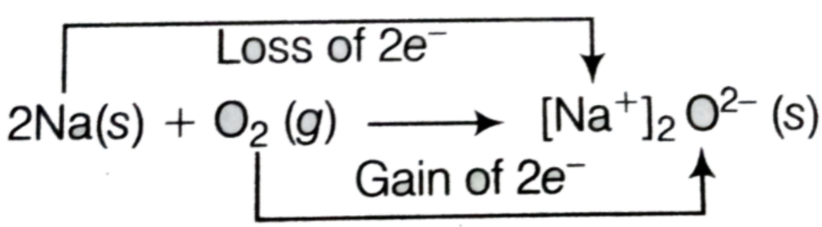
As we know, the reactions,

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
$4Na(s) + O_2(g) \rightarrow 2Na_2(s)$

because each of these reactions involves the addition of oxygen or a more electronegative element to sodium, they are redox reactions.

At the same time, chlorine and oxygen are being depleted, while sodium, an electropositive element, has been added. We also know that sodium chloride and sodium oxide are ionic compounds and are perhaps better expressed as $Na^+Cl^-(s)$ and $Na_2^+O_{2,}$ respectively, based on our understanding of chemical bonding(s). The development of charges on the species created leads us to recast the preceding reaction:





Each of the above processes can be thought of as two different phases, one involving electron loss and the other involving electron gain. We can expound on one of these, say sodium chloride production, as an example.

$2Na(s) \rightarrow 2Na^+(g) + 2e^-$
$Cl_2 + 2e^- \rightarrow 2Cl^-(g)$

Each of the preceding steps is referred to as a half-reaction since it clearly demonstrates the involvement of electrons. The total reaction is equal to the sum of the half-reactions:

$2Na(s) + Cl^2(g) \rightarrow 2Na + Cl^-(s)$ or $2NaCl(s)$

The above reactions indicate that 50% of the reactions involving electron loss are oxidation reactions. Similarly, half-reactions involving electron gain are referred to as reduction reactions.

The basic operations of life, such as photosynthesis, respiration, combustion, and corrosion or rusting, are redox reactions.

Q2. On the basis of standard electrode potential values, suggest which of the following reactions would take place? (Consult the book for E° value).

(i) $Cu + Zn^{2+} \longrightarrow Cu^{2+} + Zn$
(ii) $Mg + Fe^{2+} \longrightarrow Mg^{2+} + Fe$
(iii) $Br_2 + 2Cl^- \longrightarrow Cl_2 + 2Br^-$
(iv) $Fe + Cd^{2+} \longrightarrow Cd + Fe^{2+}$

Answer:

The cell's net cell EMF determines whether or not a reaction will occur. The formula is as follows:

$$E^o_{cell} = E^o_{cathode} - E^o_{anode}$$

Option (ii) clearly shows that the reaction can occur because Mg has a lower Ecirc cell value. As a result, Mg is oxidised by losing an electron, whereas iron is reduced by gaining one.

$$Mg + Fe^{2+} \rightarrow Mg^{2+} + Fe$$

We can state that Fe goes through reduction and Mg goes through oxidation.

Taking the E° values,

$E^o_{cathode}$ = -0.44V

E^o_{anode} = -2.36V

E^o_{cell} = -0.44 - (-2.36)V

E^o_{cell} = +1.92V

Q3. Why does fluorine not show a disproportionation reaction?

Answer:

Disproportionation is a redox reaction in which one intermediate oxidation state component transforms into two higher and lower oxidation state compounds.

The element must be in at least three oxidation states for such a redox reaction to occur. As a result, that element is in the intermediate state during the disproportionation reaction and can transition to both higher and lower oxidation states.

Fluorine is the most electronegative and oxidising element of all the halogens, and it is also the smallest.

It doesn't have a positive oxidation state (only one) and doesn't go through the disproportionation reaction.

Q4. Find out the oxidation number of chlorine in the following compounds and arrange them in increasing order of oxidation number of chlorine.

$NaClO_4$, $NaClO_3$, $NaClO$, $KClO_2$, Cl_2O_7, ClO_3, Cl_2O, $NaCl$, Cl_2, ClO_2.

Which oxidation state is not present in any of the above compounds?

Answer:

Considering oxidation number of chlorine is ‘x’, then 1 + x + 4 * (-2) = 0
∴ x - 7 = 0
∴ x = +7

$Na^{(+1)}Cl^{(+7)}O_4^{(-2)}$

Here, oxidation number of chlorine is +7.

Following the above method,

$Na^{(+1)}Cl^{(+1)}O_3^{(-2)}$	Oxidation number of chlorine = +5
$Na^{(+1)}Cl^{(+1)}O^{(-2)}$	Oxidation number of chlorine = +1
$K^{(+1)}Cl^{(+3)}O_2^{(-2)}$	Oxidation number of chlorine = +3
$Cl_2^{(+7)}O_7^{(-2)}$	Oxidation number of chlorine = +7
$Cl^{(+6)}O_3^{(-2)}$	Oxidation number of chlorine = +6
$Cl_2^{(+1)}O^{(-2)}$	Oxidation number of chlorine = +1
$Na^{(+1)}Cl^{(-1)}$	Oxidation number of chlorine = -1
Cl_2	Oxidation number of chlorine = 0
$Cl^{(+4)}O_2^{(-2)}$	Oxidation number of chlorine = +4

The following compounds are listed in increasing order of chlorine oxidation number:

$NaCl$, Cl_2, Cl_2O, $NaClO$, $KClO_2$, ClO_2, $NaClO_3$, ClO_3, Cl_2O_7, $NaClO_4$.

Q5. Which method can be used to find out the strength of the reductant/oxidant in a solution? Explain with an example.

Answer:

When a reductant (reducing agent) or oxidant (oxidising agent) is linked in a solution using a cell, the relative electrode potential can be measured. The element under discussion can be used as an electrode in a conventional cell with a known electrode potential. The electrode of the given species works as a reductant if it is positive and an oxidant if it is negative.

For example, we want to test Fe^{3+}/Fe with a Standard Hydrogen electrode (SHE). For Fe and H, the half-cell reaction is as follows:

$H^+ + e^- \rightarrow H_2 E^o = 0.0V$

$Fe^{3+} + e^- \rightarrow Fe^{2+} E^o = 0.77$

With SHE, any element that needs to be evaluated is used as an electrode. The element's potential is defined as the quantity of emf it generates in the cell.

$E^o_{cell} = E^o_{cathode} - E^o_{anode}$

$E^o_{cell} = 0 - E^o_{anode}$

$E^o_{cell} = 0 - 0.77$

$E^o_{cell} = -0.77$

In comparison to hydrogen, the Fe^{3+} ion has a higher tendency to undergo reduction. As a result, the previously assumed Fe anode configuration can be reversed, and the strength of Fe as a reductant can be established. As a result, the oxidant's strength can be determined.