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Class: 11
Subject: Chemistry
Topic: Redox Reactions
No. of Questions: 20

Q1. Justify that the reaction
$2Cu_2O_{(s)}+Cu_2S_{(s)} \rightarrow 6Cu_{(s)}+SO_{2(g)}$ is a redox reaction.

Sol. Writing the oxidation number of each atom above its symbol, we have

$$2\overset{+1}{Cu_2}\overset{-2}{O}_{(s)} + \overset{+1}{Cu_2}\overset{-2}{S}_{(s)} \rightarrow 6\overset{0}{Cu}_{(s)} + \overset{+4}{S}\overset{-2}{O}_{2(g)}$$

Here in the reaction, the oxidation number of copper decreases from +1 in Cu_2O or Cu_2S to 0 in copper metal, therefore copper is reduced. Further the oxidation number of S increases from -2 in Cu_2S to +4 in SO_2, therefore S is oxidised. Therefore, this reaction is a redox reaction.

Q2. Which of the following species do not show disproportionation reaction and why?
ClO^-, ClO_2^-, ClO_3^- and ClO_4^-

Sol. Out of the above species, ClO_4^- does not undergo disproportionation since in this oxoanion chlorine is already present in the highest oxidation state of +7 and hence cannot be further oxidised.

Q3 Permanganate ion reacts with bromide ion in basic medium to give manganese dioxide and bromate ion. Write the balanced chemical equation for the reaction.

Sol. $2MnO_4^-{}_{(aq)} + MnO_4^-{}_{(aq)} + H_2O_{(liq)} \rightarrow 2MnO_{2(s)} + BrO_3^-{}_{(aq)} + 2OH^-{}_{(aq)}$

Q4. Suggest a scheme of classification of the following redox reaction:

$2NO_{2(g)} + 2OH^-{}_{(aq)} \rightarrow NO^-{}_{2(aq)} + NO^-{}_{3(aq)} + H_2O_{(liq)}$

Sol. This is a disproportionation reaction since here the oxidation state of nitrogen decreases from +4 in NO_2 to +3 in NO_2^- ion, as well as increases from +4 in NO_2 to +5 in NO_3^- ion.

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Q5. In which compounds is the oxidation number of Oxygen -1 and +2?

Sol. The oxidation number in peroxides (Na_2O_2) is -1 and in OF_2 its oxidation number is +2.

Q6. Is it possible to store Copper Sulphate Solution in a Zinc vessel?

Sol. We cannot place $CuSO_4$ solution in a zinc vessel, if the following redox reaction occurs:
$Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$ or $Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$
By convention, the cell may be represented as $Zn|Zn^{2+}\|Cu^{2+}|Cu$. Therefore,
$E^{o}_{cell}=E^{o}Cu^{2+}, Cu; E^{o}Zn^{2+}, Zn=0.34-(-0.76)= +1.10V$. Since, EMF comes out to be positive, therefore, $CuSO_4$ reacts with Zinc. In other words, $CuSO_4$ solution cannot be stored in a Zinc vessel.

Q7. Calculate the oxidation number of S in $Na_2S_4O_6$.

Sol. Let the oxidation number of S be X. Then

+1	X	-2
Na_2	S_4	O_6
(+1 x 2)	(X x 4)	(-2 x 6)

2 + 4X − 12 = 0 or 4X = 10 or X = 10/4 which is 2.5

Q8. Is the sum of oxidation numbers of all atoms in an ion is equal to 0?

Sol. No, the sum of oxidation number of all atoms in an ion is equal to the charge on the ion.

Q9. Can the oxidation number of an element be zero?

Sol. Yes, the oxidation number of an element can be zero. For example, the oxidation number of carbon in CH_2Cl_2 is zero.

Q10 The electrode potential of four metallic elements (A, B, C and D) are +0.80V, -0.76V, +0.12V and +0.34V respectively. Arrange them in order of decreasing electropositive character.

Sol. Higher the electrode potential (E^0) lower is the tendency of the metal to lose electrons and hence lower is the electropositive character of the metal.

Q11. Define oxidation in terms of oxidation numbers.

Sol. Oxidation may be defined as a chemical change in which there occurs an increase in the oxidation number of an atom or atoms.

Q12. Justify that the reaction $2Na_{(s)} + H_{2(g)} \rightarrow 2NaH_{(s)}$ is a redox reaction.

Sol. Since NaH is an ionic compound, it may be represented as $Na^{+} H^{-}{}_{(s)}$

$2Na_{(s)} + H_{2(g)} \rightarrow 2Na^{+} H^{-}$

This reaction can be split into following two half reactions:

$Na_{(s)} \rightarrow Na^{+} + e^{-}$ (oxidation) (loss of electrons)

$H_{2(s)} + 2e^{-} \rightarrow 2H^{-}$ (reduction) (gain of electrons)

Overall redox reaction,

$2Na_{(s)} + H_{2(g)} \rightarrow 2Na^{+} H^{-}{}_{(s)}$

Q13. The half-cell reactions with their Oxidation potentials are:

$Pb_{(s)} \rightarrow Pb^{2+}_{(aq)} + 2e^{-}; E^{o}_{oxi} = +0.13V$
$Ag_{(s)} \rightarrow Ag^{+}_{(aq)} + e^{-};_{oxi} = 0.80V$
Write the cell reaction and calculate its EMF.

Sol. Re-write the two equations in the reduction form. Thus
$Pb^{2+}{}_{(aq)} + 2e^{-} \longrightarrow Pb_{(s)}$ $E^{o} = -0.13V$ ....(i)
$Ag^{+}{}_{(aq)} + e^{-} \longrightarrow Ag_{(s)}$ $E^{o} = +0.80V$ ....(ii)
To obtain equation for the cell reactions, multiply Eq. (ii) with 2 and subtract Eq. (i) from it; We have
$Pb_{(s)} + 2Ag^{+}{}_{(aq)} \longrightarrow Pb^{2+}{}_{(aq)} + 2Ag_{(s)}$
$E^{o}{}_{cell} = 2(+0.80) - (-0.13) = 1.6 + 0.13 = 1.73$ V

Q14. What is Standard Electrode Potential?

Sol. Electrode potential are generally measured under standard conditions in 1 mol L^{-1} and a temperature of 298 K and are called Standard Electrode potential denoted by E^{o}.

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Q15. Define Reduction in terms of oxidation numbers.

Sol. Reduction may be defined as a chemical change in which there occurs a decrease in the oxidation number of an atom or atoms.

Q16. A solution of silver nitrate was stirred with iron rod. Will it cause any change in the concentration of silver and nitrate ions?

Sol. Since E° of Fe^{2+}/Fe(-0.44V) is lower than that of Ag^{+}/Ag(+0.80 V) electrode therefore, Ag^{+} gets reduced and Fe gets oxidized. As a result, concentration of Ag+ ions decreases while that of NO_3^{-}.

$2Ag^{+}_{(aq)} + Fe_{(s)} \longrightarrow 2Ag_{(s)} + Fe^{2+}_{(aq)}$ ions remain unchanged.

Q17. The standard electrode potentials of a few metals are given below:

AI(-0.66V), Cu(+0.34V), Li(-3.05V), Ag(+0.80V) and Zn(-0.76V)
Which of these will behave as the strongest oxidizing agent and which as the strongest reducing agent?

Sol. Li is the strongest reducing agent while Ag+ is the strongest oxidizing agent.

Q18. What are the two ways by which a Redox equation is balanced?

Sol. The two methods of balancing the redox equation are:
a) Oxidation number method
b) Ion electron method or half equation method.

Q19. Choose the strongest oxidizing agent among F_2, Br_2, I_2, Cl_2.

Sol. F_2

Q20. Why is it not possible to balance the following equation?
$Cr_2O7^{2-} + Fe^{3+} + H^{+} \rightarrow Cr^{3+} + Fe^{2+} + H_2O$

Sol. In this reaction the oxidation state of Chromium is decreasing from +6 to +3 while that of Iron is also decreasing from +3 to +2. This is not a redox reaction and is not written because a substance cannot get reduced until some other sentence gets oxidized.