Important MCQs on Redox Reactions with Explanation

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Q.1. Which of the following is not an example of redox reaction?

(a) $CuO + H_2 \rightarrow Cu + H_2O$

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

(c) $2K + F_2 \rightarrow 2KF$

(d) $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$

Ans.- (d) $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$

Explanation- In option (d), there occur no change in oxidation number of any element. Hence, it is not an example of redox reaction

$$\overset{+2\ -1}{BaCl_2} + \overset{+1\ -2}{H_2SO_4} \rightarrow \overset{+2\ -2}{BaSO_4} + \overset{+1\ -1}{2HCl}$$

Other options involve change in oxidation number of the elements. Hence, they are the examples of redox reactions.

$$\overset{+2\ -2}{CuO} + \overset{0}{H_2} \rightarrow \overset{0}{Cu} + \overset{+1\ -2}{H_2O}$$

$$\overset{+3\ -2}{Fe_2O_3} + \overset{+2\ -2}{3CO} \rightarrow \overset{0}{2Fe} + \overset{+4\ -2}{3CO_2}$$

$$\overset{0}{2K} + \overset{0}{F_2} \rightarrow \overset{+1\ -1}{2KF}$$

Q.2. Standard electrode potential (E^0) of some redox couples are as under-

$E^0(Fe^{3+}/ Fe^{2+}) = +0.77$ V $E^0(Ag^+/Ag) = +0.80$ V

$E^0(Cu^{2+}/Cu) = +0.34$ V $E^0(I_2/I^-) = +0.54$ V

Which of the following is the strongest oxidizing agent.-

(a) Fe^{3+} (b) I_2 (s) (c) Cu^{2+} (d) Ag^+

Ans.- (d) Ag^+

Explanation-The greater the positive the value of standard electrode potential (E^0), more is the tendency of the species to get reduced and therefore, stronger is the oxidizing behavior of the species. On this basis Ag^+ will be the strongest oxidizing agent.

Q.3. Standard electrode potential (E^0) of some redox couples are as under-

$E^0(Br_2/Br^-) = +1.90$ V $E^0(Ag^+/Ag) = +0.80$ V

$E^0(Cu^{2+}/Cu) = +0.34$ V $E^0(I_2/I^-) = +0.54$ V

On the basis of these values, choose the correct option-

(a) Cu will reduce Br^- (b) Cu will reduce Ag

(c) Cu will reduce I^- (d) Cu will reduce Br_2

Ans.- (d) Cu will reduce Br_2

Explanation-The greater the positive the value of standard electrode potential (E^0), more is the tendency of the species to get reduced and therefore, stronger is the oxidizing behavior of the species. On this basis Br_2 will be the strongest oxidizing agent amongst and it will gain electron from Cu (reducing agent) to reduce into Br^-.

Q.4. Standard electrode potential (E^0) of some redox couples are as under-

$E^0(Fe^{3+}/ Fe^{2+}) = +0.77$ V $E^0(Ag^+/Ag) = +0.80$ V

$E^0(Cu^{2+}/Cu) = +0.34$ V $E^0(I_2/I^-) = +0.54$ V

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Using these data find out the pair between which redox reaction is not feasible.

(a) Fe^{3+} and I^- (b) Ag^+ and Cu

(c) Fe^{3+} and Cu (d) Ag and Fe^{3+}

Ans.- (d) Ag and Fe^{3+}

Explanation-The greater the positive the value of standard electrode potential (E^0), more is the tendency of the electrode to get reduced and therefore, stronger is the oxidizing behavior of the electrode. Here, $E^0(Ag^+/Ag)$ has more positive value of standard electrode potential as compared to $E^0(Fe^{3+}/ Fe^{2+})$. Thus, Ag is poor reducing agent than Fe^{2+} ion. Thus, Ag cannot reduce Fe^{3+} ions.

Q.5. Thiosulphate reacts differently with iodine and bromine in the reactions given below-

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

$$2S_2O_3^{2-} + 2Br_2 + 5H_2O \rightarrow 2SO_4^{2-} + 2Br^- + 10\ H^+$$

Which of the following statements justifies the above dual behaviour of thiosulphate?

(a) Bromine is a stronger oxidant than iodine

(b) Bromine is a weaker oxidant than iodine.

(c) Thiosulphate undergoes oxidation by bromine and reduction by iodine in these reactions.

(d) Bromine undergoes oxidation and iodine undergoes reduction in these reactions.

Ans.- (a) Bromine is a stronger oxidant than iodine

Explanation-Bromine having more positive value of standard electrode potential [$E^0(Br_2/Br^-)$ = +1.90 V] is a stronger oxidant than iodine having lesser positive value of standard electrode potential [$E^0(I_2/I^-)$ = +0.54 V]. Thus, bromine oxidizes thiosulphate ion from +2 oxidation state of sulphur to +7 oxidation state (SO_4^{2-}) while, iodine oxidizes thiosulphate ion from +2 oxidation state of sulphur to +5 oxidation state ($S_4O_6^{2-}$).

$$\overset{+2}{2S_2O_3^{2-}} + I_2 \rightarrow \underset{\text{Dithionate ion}}{\overset{+5}{S_4O_6^{2-}}} + 2I^-$$

$$\overset{+2}{2S_2O_3^{2-}} + 2Br_2 + 5H_2O \rightarrow \overset{+6}{2SO_4^{2-}} + 2Br^- + 10\ H^+$$

Q.6. The oxidation number of an element in a compound is evaluated on the basis of certain rules. Which of the following rules is not correct in this respect?

(a) The oxidation number of hydrogen is always +1.

(b) The algebraic sum of all the oxidation numbers in a compound is zero.

(c) An element in the free or the uncombined state bears oxidation number zero.

(d) In all its compounds, the oxidation number of fluorine is – 1.

Ans.- (a) The oxidation number of hydrogen is always +1.

Explanation- The oxidation number of hydrogen is usually +1. In metal hydridres (NaH, CaH_2, AlH_3 etc.) it is -1 and in uncombined state (H_2) it is zero.

Q.7. In which of the following compounds, an element exhibits two different oxidation states.

(a) NH_2OH (b) NH_4NO_3

(c) N_2H_4 (d) N_3H

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Ans.- (b) NH_4NO_3

Explanation- In NH_4NO_3 the oxidation number of two nitrogen atoms are -3 and +5 respectively.

$$\overset{-3}{N}H_4\overset{+5}{N}O_3$$

Q.8. Which of the following arrangements represent increasing oxidation number of the central atom?

(a) CrO_2^-, ClO_3^-, CrO_4^{2-}, MnO_4^-
(b) ClO_3^-, CrO_4^{2-}, MnO_4^-, CrO_2^-
(c) CrO_2^-, ClO_3^-, MnO_4^-, CrO_4^{2-}
(d) CrO_4^{2-}, MnO_4^-, CrO_2^-, ClO_3^-

Ans.- (a) CrO_2^-, ClO_3^-, CrO_4^{2-}, MnO_4^-

Explanation- The oxidation number of the central atom in given species are as under-

$\overset{+3}{Cr}O_2^-$	$\overset{+5}{Cl}O_3^-$	$\overset{+6}{Cr}O_4^{2-}$	$\overset{+7}{Mn}O_4^-$

Q.9. Which of the following is a disproportionation reaction-

(a) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
(b) $CH_4 + 4Cl_2 \rightarrow CCl_4 + 4HCl$
(c) $2F_2 + 2OH^- \rightarrow 2F^- + OF_2 + H_2O$
(d) $2NO_2 + 2OH^- \rightarrow NO_2^- + NO_3^- + H_2O$

Ans.- (d) $2NO_2 + 2OH^- \rightarrow NO_2^- + NO_3^- + H_2O$

Explanation- Such reactions in which the oxidation number of an element both increases and decreases simultaneously are called disproportionation reactions.

$$\overset{+4}{N}O_2 + 2OH^- \rightarrow \overset{+3}{N}O_2^- + \overset{+5}{N}O_3^- + H_2O$$

Q.10. Which of the following elements does not show disproportionation tendency?

(a) Cl (b) Br (c) F (d) I

Ans.- (c) F

Explanation- In combined state fluorine shows only the oxidation state of -1. Being the most electronegative element, it does not exhibit positive oxidation state and therefore, it does not show disproportionation tendency.

Q.11. Consider following reaction-

$$P_4 + 3OH^- + 3H_2O \rightarrow PH_3 + 3H_2PO_2^-$$

Identify the correct statement-

(a) Phosphorus is undergoing reduction only
(b) Phosphorus is undergoing oxidation only
(c) Phosphorus is undergoing oxidation as well as reduction
(d) Hydrogen is undergoing neither oxidation nor reduction

Ans.- (c) Phosphorus is undergoing oxidation as well as reduction

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Explanation-In given reaction, the oxidation state of phosphorus both increases and decreases.

$$\overset{0}{P_4} + 3OH^- + 3H_2O \rightarrow \overset{-3}{PH_3} + 3H_2\overset{+1}{P}O_2^-$$

Thus, it is the example of disproportionation reaction where, phosphorus is undergoing both oxidation as well as reduction.

Q.12. In which of the following processes, neither oxidation nor reduction occur-

(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

Ans.-(a) Formation of slaked lime from quick lime

Explanation-Formation of slaked lime from quick lime does not involve change in oxidation number of any element.

$$\overset{+2}{CaO} + H_2O \rightarrow \overset{+2}{Ca(OH)_2}$$

Hence, it is not a redox reaction i.e. neither oxidation nor reduction occurs in this reaction. Mercuric oxide (red coloured solid) on heating gives mercury and releases oxygen. It involves change in oxidation number and therefore, it is a redox reaction.

$$\overset{+2\ -2}{2HgO(s)} \rightarrow \overset{0}{2Hg(l)} + \overset{0}{O_2(g)}$$

Formation of manganese chloride from manganese oxide involves change in oxidation number of Mn from +4 to +2.

$$\overset{+4}{MnO_2} \rightarrow \overset{+2}{MnCl_2}$$

Formation of zinc from zinc blende also involves change in oxidation number of Zn from +2 to 0.

$$\overset{+2}{ZnS} \rightarrow \overset{0}{Zn}$$

Q.13. The oxidation number of chlorine in Cl_2O_7 is-

(a) + 7 (b) + 5 (c) + 3 (d) - 7

Answer-(a) + 7

Explanation-

Let, oxidation number of chlorine in Cl_2O_7 is x. then,

$$2x + 7 \times (-2) = 0$$
$$\therefore \quad x = +7$$

Q.14. Which of the following is the example of auto-oxidation-

(a) Formation of H_2O by the decomposition of H_2O_2.
(b) Formation of H_2O_2 by the oxidation of H_2O.
(c) Both (a) and (b)
(d) None of the above

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

Explanation-Self-oxidation of any substance in presence of oxygen is called auto-oxidation of the substance. This term is usually used to describe the degradation of organic compounds in air (as a source of oxygen) to produce hydroperoxides and cyclic organic peroxides which react further to form other products. This process is responsible for aging, spoilage of foods, degradation of petrochemicals etc. Auto-oxidation may occur in a situation where it is not desired. For example, destructive cracking of the rubber of automobile tires and rancidification.

Water also gets oxidized to hydrogen peroxide in presence of air and it is the example of auto-oxidation. Here, oxidation number of oxygen from increases from -2 to -1.

Formation of H_2O by the decomposition of H_2O_2 involves decrease of oxidation number of oxygen from -1 to -2. Hence, it is not the example of oxidation, but it is the example of reduction.

Q.15.-If equal volumes of 1M $KMnO_4$ and 1M $K_2Cr_2O_7$ solutions are allowed to oxidize Fe^{2+} in acidic medium. The amount of Fe^{2+} oxidized will be-

(a) More with $KMnO_4$ (b) More with $K_2Cr_2O_7$

(c) Equal with both the oxidizing agents (d) Cannot be determined

Ans.-(b) More with $K_2Cr_2O_7$

Explanation-In presence of H_2SO_4 (acidic medium), $KMnO_4$ changes to $MnSO_4$, where, the oxidation number of Mn changes from +7 to +2 while, $K_2Cr_2O_7$ changes to $Cr_2(SO_4)_3$, where, the oxidation number of Cr changes from +6 to +3.

$$\overset{+7}{KMnO_4} \rightarrow \overset{+2}{MnSO_4} \qquad \text{Total change in Oxidation No.} = 5$$

$$\overset{+6}{K_2Cr_2O_7} \rightarrow \overset{+3}{Cr_2(SO_4)_3} \qquad \text{Total change in Oxidation No.} = 6$$

Thus, the total change in the oxidation number i.e. n factor is more for $K_2Cr_2O_7$. Hence, for same volume, 1M $K_2Cr_2O_7$ contains more number of gram equivalent of oxidizing agent as compared to 1M $KMnO_4$. Thus, the amount of Fe^{2+}oxidized will be more for $K_2Cr_2O_7$ as compared to that for $KMnO_4$.

Q. 16.- The number of moles of $KMnO_4$ reduced by one mole of KI in alkaline medium is-

(a) 1 (b) 2 (c) 5 (d) 1/5

Answer-(b) 2

Explanation-In alkaline medium, KI reduces $KMnO_4$ as per the following reaction-

$$2KMnO_4 + H_2O \rightarrow 2KOH + 2MnO_2 + 3[O]$$

$$KI + 3[O] \rightarrow KIO_3$$

Hence, the overall reaction is-

$$KI + 2KMnO_4 + H_2O \rightarrow KIO_3 + 2\,KOH + 2\,MnO_2$$

Thus, in alkaline medium, one mole of KI will reduced two moles of $KMnO_4$.

Q.17.-The oxidation number of Xe in $BaXeO_6$ is-

(a) 8 (b) 6 (c) 4 (d) 10

Ans.-(a) 8

Explanation-Let, oxidation number of Xe in $BaXeO_6$ is x. then,

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$+2 + x + 6 \times (-2) = 0$

$\therefore \quad x = +10$

However, oxidation state of +10 is not possible for Xe. Hence, in this case the oxidation state of Xe will be equal to its maximum possible oxidation state i.e. +8.

Q.18. $KMnO_4$ reacts with oxalic acid as per the following 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 $C_2H_2O_4$

Answer-(d) 50 ml of 0.1 M $C_2H_2O_4$

Explanation-

$2MnO_4^- + 5C_2O_4^{2-} + 16H+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$

Here, 2 moles of MnO_4^- is equivalent to 5 moles of $C_2O_4^{2-}$

20 ml of 0.1 M $KMnO_4$ = 2 mili-moles of $KMnO_4$ $\equiv$ 5 mili-moles of $C_2O_4^{2-}$

It corresponds to option (d) i.e. 50 ml of 0.1 M $C_2H_2O_4$.

Q.19. One mole of N_2H_4 loses ten moles of electrons to form a new compound A. Assuming that all the nitrogen atoms appear in the new compound, what will be the oxidation state of nitrogen in A? (There is no change in the oxidation state of hydrogen.)

(a) -1 (b) -3 (c) +3 (d) +5

Answer-(c) +3

Explanation- Let, per atom oxidation number of nitrogen in N_2H_4 is x. Then,

$2x + 4(+1) = 0$

$\therefore \quad 2x = -4$

or $\quad x = -2$

Thus, total oxidation number of nitrogen atoms in one mole of $N_2H_4 = 2 \times -2 = -4$

Total change in oxidation number of nitrogen after losing 10 mol of electrons (considering no change in oxidation number of hydrogen atoms)

$-4 - (-10) = +6$

Thus, oxidation number of 2 nitrogen atoms in new compound A will be +6. Hence, oxidation number of each nitrogen atom in new compound A will be +3.

Q. 20. Which of the following defines n-factor of an oxidant or reductant ?

(a) Number of moles of electrons lost or gained by one mole of reductant or oxidant

(b) Total change in oxidation number of one mole of reductant or oxidant

(c) Both the above

(d) None of the above

Ans.-(c) Both the above

Explanation- n-factor of an oxidant or reductant is defined as number of moles of electrons lost or gained by one mole of reductant or oxidant during a balanced redox reaction. It is also defined as the total change in oxidation number of one mole of reductant or oxidant during the redox reaction.

For example, in presence of H_2SO_4 (acidic medium), $KMnO_4$ changes to $MnSO_4$, where, the oxidation number of Mn changes from +7 to +2 while, $K_2Cr_2O_7$ changes to $Cr_2(SO_4)_3$, where, the oxidation number of Cr changes from +6 to +3.

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$\overset{+7}{KMnO_4} \rightarrow \overset{+2}{MnSO_4}$ Total change in Oxidation No. = 5

$\overset{+6}{K_2Cr_2O_7} \rightarrow \overset{+3}{Cr_2(SO_4)_3}$ Total change in Oxidation No. = 6

Here, n-factor of $KMnO_4$ in acidic medium will be 5 while, that of $K_2Cr_2O_7$ will be 6.

Q.21. The standard reduction electrode potentials of four metals A, B, C, and D are -0.250 V, -0.150 V, -0.540 V and +0.340 V respectively. The metal that will displaces A from its aqueous solution is-

(a) A (b) B (c) C (d) D

Answer-(c) C

Explanation-Here, metal C has more negative value of electrode potential (-0.540 V) as compared to metal A (-0.250 V). Hence, C will be more reactive than A and it will displace metal A from its salt solution.

Q.22. Which of the following change involves neither oxidation nor reduction ?

(a) $VO^{2+} \rightarrow V_2O_3$

(b) $Na \rightarrow Na^+$

(c) $CrO_4^{2-} \rightarrow Cr_2O_7^{2-}$

(d) $Zn^{2+} \rightarrow Zn$

Answer-(c) $CrO_4^{2-} \rightarrow Cr_2O_7^{2-}$

Explanation-

In $VO^{2+} \rightarrow V_2O_3$, V is reduced from +4 to +3 oxidation state.

In $Na \rightarrow Na^+$, Na is oxidised from 0 to +1 oxidation state.

In $CrO_4^{2-} \rightarrow Cr_2O_7^{2-}$, oxidation state of Cr remains unchanged i.e. +6.

In $Zn^{+2} \rightarrow Zn$, Zn is reduced from +2 to 0 oxidation state.

Q.23. How many millilitres of 0.5 M H_2SO_4 are needed to dissolve 0.5 g of copper(II)carbonate ?

(a) 6.01 (b) 4.5 (c) 8.1 (d) 11.1

Answer- (c) 8.1

Explanation-The volume will be calculated by the the formula of equivalence-

$N_1V_1 = N_2V_2$

Where, N_1 = Normality of H_2SO_4

$= 0.5 \times 2 = 1$ N

V_1 = Volume of H_2SO_4 = ?

N_2 = Normality of copper (II) carbonate

$= (0.5\times2)/(123.5 \times 1)$ N

(Molar mass of copper(II) carbonate = 123.5 g)

V_2 = Volume of copper (II) carbonate = 1000 ml

Here,

$1 \times V_1 = (0.5\times2)/(123.5)\times1000$

$\therefore \quad V_1 = 8.09$ ml

$= 8.1$ ml (approx.)

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Q.24. The oxidation state of Cr in $Cr(CO)_6$ is-

(a) 0 (b) 2 (c) 3 (d) 6

Ans.- (a) 0

Explanation- CO (carbonyl group) is a neutral ligand, hence, oxidation state of Cr in $Cr(CO)_6$ is zero.

Q.25. Which of the following processes does not involve oxidation of iron ?

(a) Formation of $Fe(CO)_5$ from Fe.

(b) Reaction of iron with steam at high temperature to liberate H_2

(c) Rusting of iron sheets.

(d) Decolourisation of blue $CuSO_4$ solution by iron.

Ans.-(a) Formation of $Fe(CO)_5$ from Fe.

Explanation-

$$Fe + 5CO \rightarrow Fe(CO)_5$$

Here, in both Fe and $Fe(CO)_5$, the oxidation state of iron is zero. Thus, oxidation number of Fe remains unchanged during the process.

$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$ (Oxidation number of Fe changes from 0 to +3)

steam

$Fe + xH_2O \rightarrow Fe_2O_3.xH_2O$ (Oxidation number of Fe changes from 0 to +3)

Rust

$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$

(Oxidation number of Fe changes from 0 to +2)

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