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| **INDIAN SCHOOL AL WADI AL KABIR** | | |
| **CLASS: XI** | **DEPARTMENT: SCIENCE (2024-25)**<br>**SUBJECT: CHEMISTRY** | **DATE: 28/10/2024** |
| **WORKSHEET NO: 5 WITH ANSWERS** | **TOPIC REDOX REACTIONS** | **NOTE: A4 FILE FORMAT** |
| **CLASS & SEC:** | **NAME OF THE STUDENT:** | **ROLL NO.** |

**Objective Type Questions**

1. KMnO4 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 $C_2H_2O_4$

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

3. The oxidation number of Xe in $BaXeO_6$ is
(a) 8
(b) 6
(c) 4
(d) 10

4. Oxidation process involves
(a) Increase in oxidation number
(b) Decrease in oxidation number
(c) No change in oxidation number
(d) None of the above

5. The oxidation number of Fe in $K_4$ [Fe $(CN)_6$] is
(a) 3
(b) 4
(c) 2
(d) 1

6. 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

7. The more positive the value of $E^{o}$, the greater is the tendency of the species to get reduced. Using the standard electrode potential of redox couples given below to find out which of the following is the strongest oxidising agent?

$E^{o}$ Values: $Fe^{3+}/Fe^{2+}$ = +0.77; $I_{2(s)}/I^{-}$ = +0.54; $Cu^{2+}/Cu$ = + 0.34; $Ag^{+}/Ag$ = + 0.80V

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

8. Using the standard electrode potential, find out the pair between which redox reaction is not feasible.

Values: $Fe^{3+}/ Fe^{2+}$ = +0.77; $I_2/I^-$= +0.54; $Cu^{2+}/ Cu$ = 0.34; $Ag^{+} /Ag$ = + 0.80 V

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

**Questions 9- 10 are *Assertion Reason* type questions**

a. If both *Assertion* and *Reason* are correct and *Reason* is the correct explanation of *Assertion*.
b. If both *Assertion* and *Reason* are correct but *Reason* is not the correct explanation of *Assertion*.
c. If *Assertion* is correct and *Reason* is wrong.
d. If *Assertion* is wrong and *Reason* is correct.

9. Assertion (A): Among halogens fluorine is the best oxidant.
Reason (R): Fluorine is the most electronegative atom.

10. 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.

**2 Marks questions**

11. Assign oxidation numbers to the underlined elements in each of the following species:
(a) $NaH_2\underline{P}O_4$ (b) $NaH\underline{S}O_4$ (c) $H_4\underline{P_2}O_7$ (d) $K_2\underline{Mn}O_4$

12. Write the formulae for the following compounds:

(a) Mercury(II) chloride (b) Nickel(II) sulphate (c) Tin(IV) oxide (d) Thallium(I) sulphate

13. Identify the substance oxidised, reduced, oxidising agent and reducing agent for each of the following reactions:

(a) $2AgBr\ (s) + C_6H_6O_{2(aq)} \rightarrow 2Ag_{(s)} + 2HBr_{(aq)} + C_6H_4O_{2(aq)}$

(b) $HCHO_{(l)} + 2[Ag(NH_3)_2]^{+}_{(aq)} + 3OH^{-}_{(aq)} \rightarrow 2Ag_{(s)} + HCOO^{-}_{(aq)} + 4NH_{3(aq)} + 2H_2O_{(l)}$

14. Balance the following redox reactions:

(a) $MnO_4^{-}{}_{(aq)} + I^{-}{}_{(aq)} \rightarrow MnO_{2\,(s)} + I_{2(s)}$ (in basic medium)

(b) $MnO_4^{-}{}_{(aq)} + SO_{2(g)} \rightarrow Mn^{2+}_{(aq)} + HSO_4^{-}{}_{(aq)}$ (in acidic solution)

15. Consider the elements:
Cs, Ne, I and F
(a) Identify the element that exhibits only negative oxidation state.
(b) Identify the element that exhibits only positive oxidation state.
(c) Identify the element that exhibits both positive and negative oxidation states.
(d) Identify the element which exhibits neither the negative nor does the positive oxidation state

16. Given the standard electrode potentials,
$K^+/K$ = –2.93V, $Ag^+/Ag$ = 0.80V, $Hg^{2+}/Hg$ = 0.79V, $Mg^{2+}/Mg$ = –2.37V. $Cr^{3+}/Cr$ = –0.74V
Arrange these metals in their increasing order of reducing power.

17. Depict the galvanic cell in which the reaction $Zn_{(s)} + 2Ag^+_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + 2Ag_{(s)}$ takes place, further show:

(i) which of the electrode is negatively charged,

(ii) individual reaction at each electrode.

**3 Marks Questions**

18.(a) Balance the following redox reaction in acidic medium.
$NO_2^- + MnO_4^- \longrightarrow Mn^{2+} + NO_3^-$
(b) Using Stock notation represent the following compounds:
(i) $HAuCl_4$
(ii) $SnO_2$

19. Explain why $3Fe_3O_{4(s)} + 8Al_{(s)} \rightarrow 9Fe_{(s)} + 4Al_2O_{3(g)}$ is an oxidation reaction?

20. (a) What is the electrode potential of a standard hydrogen electrode?

(b) An electrochemical cell is constituted by combining Al electrode ($E^0$=−1.66V) and Cu electrode ($E^0$=+0.34V). Which of these electrodes will work as a cathode and why?

21. Write the following redox reactions in the oxidation and reduction half-reaction reactions in the oxidation and reduction half-reactions.
$2K_{(s)} + Cl_{2(g)} \rightarrow 2KCl_{(s)}$
22. Balance the following equation:
$MnO_4^- + C_2O_4^{2-} \longrightarrow Mn^{2+} + CO_2$ ( in acidic medium)

**Case study-based Questions (4 marks)**

23. 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 + 2The 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 16elements 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.

(i) Negative $E^o$ 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

(ii) The oxidation number of P in $PO_4^{3-}$ is

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

(iii) Which of the following is not an example of redox reaction

(a) $CuO + H_2 \longrightarrow Cu + H_2O$
(b) $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$
(c) $2K + F_2 \longrightarrow 2KF$
(d) $BaCl_2 + H_2SO_4 \longrightarrow BaSO_4 + 2HCl$

(iv) What is the oxidation state of Sulphur in $Na_2SO_4$ ?

(a) -4
(b) -6
(c) +6
(d) +4

24. 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 appears 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.

(i) Highest oxidation state of Mn is present in:
(a)$KMnO_4$
(b)$K_2MnO_4$
(c)$Mn_2O_3$
(d)$MnO_2$

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

(a) Oxygen

(b) Chlorine

(c) Fluorine

(d) Bromine

(iii) The brown ring complex compound is formulated as $[Fe(H_2O)_5NO]SO_4$. What will be the oxidation state of iron in the given complex?

(a) + 2

(b) + 3

(c) + 4

(d) + 1

(iv) Redox reactions involve:

(a) Transfer of electrons between atoms or ions
(b) Breaking and forming of covalent bonds
(c) Changes in the physical state of a substance
(d) Absorption or release of heat

**5 Marks Questions**

25. On the basis of standard electrode potential values, suggest which of the following reactions would take place?

(i) $Cu + Zn^{2+} \longrightarrow Cu^{2+} + Zn$ ($E^0$ $Zn^{2+}/Zn$) = -0.76 V, ($E^0$ $Cu^{2+}/Cu$) = +0.34 V

(ii) $Mg + Fe^{2+} \longrightarrow Mg^{2+} + Fe$ ($E^0$ $Mg^{2+}/Mg$) = -2.37 V, ( $E^0$ $Fe^{2+}/Fe$) = -0.74 V

(iii) $Br_2 + 2Cl^- \longrightarrow Cl_2 + 2Br^-$ ( $E^0$ Cl-/Cl) = +1.08 V, ($E^0$ $Br^-/Br$) = +1.36 V

(iv) $Fe + Cd^{2+} \longrightarrow Cd + Fe^{2+}$ ($E^0$ $Cd^{2+}/Cd$) = -0.44 V, ($E^0$ $Fe^{2+}/Fe$) = -0.74 V

25. Find out the Oxidation number of Chlorine in the following compounds: (a) $NaClO_4$, (b) NaClO,

(c) $NaClO_3$, (d) $KClO_2$, (e) $Cl_2O_7$, (f) $ClO_3$, (g) $Cl_2O$, (h) NaCl, (i) $Cl_2$, (j) $ClO_2$

**Answers**

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|---|---|
| 1. | (d) |
| 2. | (d) |
| 3. | (d) |
| 4. | (a) |
| 5. | (c) |
| 6. | (b) |
| 7. | (d) |
| 8. | (d) |
| 9. | b. If both *Assertion* and *Reason* are correct but *Reason* is not the correct explanation of *Assertion*. |
| 10 | c. If *Assertion* is correct and *Reason* is wrong |
| 11 | (a) +5 (b) +6 (c) +5 (d) +6 |
| 12 | $Hg(II)Cl_2$(b)$Ni(II)SO_4$(c)$Sn(IV)O_2$(d)$Tl_2(I)SO_4$ |
| 13 | (a) Oxidised substance → $C_6H_6O_2$ Reduced substance → AgBr Oxidising agent → AgBr Reducing agent → $C_6H_6O_2$<br>(b) Oxidised substance → HCHO Reduced substance → $[Ag(NH_3)_2]^+$ Oxidising agent → $[Ag(NH_3)_2]^+$<br>Reducing agent → HCHO |
| 14 | (a) $6I^-_{(aq)} + 2MnO_4^-{}_{(aq)} + 4H_2O \longrightarrow 3I_{2(s)} + 2MnO_{2\,(s)} + 8OH^-_{(aq)}$<br>(b) $2MnO_4^-{}_{(aq)} + 5SO_{2(g)} + 2H_2O_{\,(l)} + H^+_{(aq)} \longrightarrow 2Mn^{2+}_{(aq)} + 5HSO_4^-{}_{(aq)}$ |

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| 15 | (a) F<br>(b) Cs<br>(c) I<br>(d) Ne |
| 16 | $Ag^+/Ag$, $Hg^{2+}/Hg$, $Cr^{3+}/Cr$, $Mg^{2+}/Mg$, $K^+/K$ |
| 17 | (i) Ag<br>(ii) $Zn(s) \rightarrow Zn^{2+}_{(aq)} + 2e^-$<br>$2Ag^+ + 2e^- \longrightarrow 2Ag$ |
| 18 | (a) $2MnO_4^- + 5NO_2^- + 6H^+ ==> 2Mn^{2+}\ 5NO_3^- + 3H_2O$<br>(b) $HAu(III)Cl_4$**.** |
| 19 | Because Al gains electrons and Fe loses electrons. |
| 20 | (a) 0 V<br>(b) Cu electrode because has $E^o$ value high indicates that have tendency to get reduced and hence it is a cathode |
| 21 | Oxidation half reaction<br>$2K(s) \rightarrow 2K^+ + 2e^-$<br>Reduction half reaction<br>$Cl_{2(g)} + 2e^- \rightarrow 2Cl^-$ |
| 22 | $2MnO_4$--$+16H^+ + 5C_2O_4^{2-} \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$. |
| 23 | (i) (b)<br>(ii) (c)<br>(iii) (d)<br>(iv) (c) |
| 24 | (i) (a)<br>(ii) (c)<br>(iii) (b)<br>(iv) (a) |
| 25 | (ii) and (iv) would take place<br>$E^0$ value +ve for these reactions |
| 26 | $NaClO_4$ - Oxidation number of Cl is +7<br>NaClO Oxidation number of Cl is +1<br>$NaClO_3$ -Oxidation number of Cl is +5<br>$KClO_2$ Oxidation number of Cl is +3<br>$ClO_3$ Oxidation number of Cl is +6<br>$Cl_2O_7$ Oxidation number of Cl is +7<br>$Cl_2O$ Oxidation number of Cl is +1<br>NaCl Oxidation number of Cl is -1<br>$Cl_2$ Oxidation number of Cl is 0<br>$ClO_2$ Oxidation number of Cl is +4 |

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