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Jars Education

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Time : 1 Hour 15 Minute

STD 11 Science Chemistry
Chapter Based Test

Total Marks : 40

SECTION A

*** Choose The Right Answer From The Given Options.[1 Marks Each]** **[6]**

1. he ratio of oxygen atom having −2 and −1 oxidation numbers in $S_2O_8^{2-}$ is _____.
   (A) 1 (B) 2 (C) 3 (D) 4
2. When $KMnO_4$ is reduced with oxalic acid in acidic solution, the oxidation number of Mn changes from.
   (A) 7 to 2 (B) 7 to 4 (C) 7 to 6 (D) 6 to 2
3. Oxidation state of Fe in Sodium Nitroprusside is:
   (A) +3 (B) +4 (C) +2 (D) +1
4. In the reaction, $2KClO_3 \rightarrow 2KCl + 3O_2$, the elements which have been oxidised and reduced respectively are:
   (A) Chlorine and oxygen. (B) Oxygen and chlorine.
   (C) Potassium and oxygen. (D) Oxygen and potassium.
5. Standard reduction frotential of X, Y, Z are -1.2v, +0.5v, -3.0v respectively, the reducing power of the metals will be:
   (A) Y > Z > X (B) Y > X > Z
   (C) Z > X > Y (D) X > Y > Z
6. Which of the following metal displacement reaction will not take place and why?
   (A) $Cu + Mg^{2+} \longrightarrow$ (B) $Mg + Cu^{2+} \longrightarrow$
   (C) $Pb + Ag^{+} \longrightarrow$ (D) $Zn + Cu^{2+} \longrightarrow$

*** Answer The Following Questions In One Sentence.[1 Marks Each]** **[5]**

7. Refer to the periodic table given in your book and now answer the following questions: Select three metals that can show disproportionation reaction.
8. What is the oxidation number of Mn in $KMnO_4$?
9. Define oxidation and reduction according to electronic concept.
10. Write redox couples involved in the reactions (i) to (iv) given in question 34.
11. Define the term redox titration.

SECTION B

*** Given Section consists of questions of 2 marks each.** **[6]**

1. Write the name of cell in which chemical energy is converted into electrical energy.

2. a. Define the term redox couple. Write the practical application of redox couple.
   b. Split $2K(s) + Cl_2(g) \longrightarrow 2KCl(s)$ into oxidation and reducton half reaction.

3. i. Balance $MnO_4^- + Fe^{2+} \longrightarrow Fe^{3+} + Mn^{2+}$ in acidic medium by ion electron method.
   ii. Given the standard electrode potentials:

$$\frac{K_+}{K} = -2.93V,$$
$$\frac{Ag^+}{Ag} = +0.80V,$$
$$\frac{Mg^{2+}}{Mg} = -2.37V$$

Arrange these metals in order of increasing reducing power.

SECTION C

*** Given Section consists of questions of 3 marks each.** **[9]**

1. Consider the reactions:
   a. $H_3PO_2(aq) + 4AgNO_3(aq) + 2H_2O(l) \rightarrow H_3PO_4(aq) + 4Ag(s) + 4HNO_3(aq)$
   b. $H_3PO_2(aq) + 2CuSO_4(aq) + 2\ H2O(l) \rightarrow H_3PO_4(aq) + 2Cu(s) + H_2SO_4(aq)$
   c. $C_6H_5CHO(l) + 2[Ag(NH_3)_2]^+(aq) + 3OH^-(aq) \rightarrow C_6H_5COO^-(aq) + 2Ag(s) + 4NH_3(aq) + 2H_2O(l)$
   d. $C_6H_5CHO(l) + 2Cu^{2+}(aq) + 5OH^-(aq) \rightarrow$ No change observed.

   What inference do you draw about the behaviour of Ag^+ and Cu^{2+} from these reactions?

2. Balance the following ionic equations.

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

3. Using the standard electrode potentials given in the Table, predict if the reaction between the following is feasible:

Reaction (Oxidised form + ne⁻ → Reduced form)	$E^\ominus$ / V
$F_2(g) + 2e^- \rightarrow 2F^-$	2.87
$Co^{3+} + e^- \rightarrow Co^{2+}$	1.81
$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	1.78
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.51
$Au^{3+} + 3e^- \rightarrow Au(s)$	1.40
$Cl_2(g) + 2e^- \rightarrow 2Cl^-$	1.36
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	1.33
$O_2(g) + 4H^+ + 4e^- \rightarrow 2H_2O$	1.23
$MnO_2(s) + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$	1.23
$Br_2 + 2e^- \rightarrow 2Br^-$	1.09
$NO_3^- + 4H^+ + 3e^- \rightarrow NO(g) + 2H_2O$	0.97
$2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$	0.92
$Ag^+ + e^- \rightarrow Ag(s)$	0.80
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.77
$O_2(g) + 2H^+ + 2e^- \rightarrow H_2O_2$	0.68
$I_2(s) + 2e^- \rightarrow 2I^-$	0.54
$Cu^+ + e^- \rightarrow Cu(s)$	0.52
$Cu^{2+} + 2e^- \rightarrow Cu(s)$	0.34
$AgCl(s) + e^- \rightarrow Ag(s) + Cl^-$	0.22
$AgBr(s) + e^- \rightarrow Ag(s) + Br^-$	0.10
$2H^+ + 2e^- \rightarrow H_2(g)$	**0.00**
$Pb^{2+} + 2e^- \rightarrow Pb(s)$	−0.13
$Sn^{2+} + 2e^- \rightarrow Sn(s)$	−0.14
$Ni^{2+} + 2e^- \rightarrow Ni(s)$	−0.25
$Fe^{2+} + 2e^- \rightarrow Fe(s)$	−0.44
$Cr^{3+} + 3e^- \rightarrow Cr(s)$	−0.74
$Zn^{2+} + 2e^- \rightarrow Zn(s)$	−0.76
$2H_2O + 2e^- \rightarrow H_2(g) + 2OH^-$	−0.83
$Al^{3+} + 3e^- \rightarrow Al(s)$	−1.66
$Mg^{2+} + 2e^- \rightarrow Mg(s)$	−2.36
$Na^+ + e^- \rightarrow Na(s)$	−2.71
$Ca^{2+} + 2e^- \rightarrow Ca(s)$	−2.87
$K^+ + e^- \rightarrow K(s)$	−2.93
$Li^+ + e^- \rightarrow Li(s)$	−3.05

Increasing strength of oxidising agent ↑ / ↓ Increasing strength of reducing agent

Ag(s) and $Fe^{3+}(aq)$.

SECTION D

*** Case study based questions** **[4]**

1. Read the passage given below and answer the following questions from 1 to 5.

   Chemistry deals with varieties of matter and change of one kind of matter into the other. Transformation of matter from one kind into another occurs through the various types of reactions. One important category of such reactions is Redox Reactions.

Originally, the term oxidation was used to describe the addition of oxygen to an element or a compound. Because of the presence of dioxygen in the atmosphere (~20%), many elements combine with it and this is the principal reason why they commonly occur on the earth in the form of their oxides. The following reactions represent oxidation processes:

$$2Mg(s) \rightarrow 2MgO\ (S)$$

$$S(s) + O_2(g) \rightarrow SO_2\ g$$

the term "oxidation" is defined as the addition of oxygen/electronegative element to a substance or removal of hydrogen/ electropositive element from a substance. In the beginning, reduction was considered as removal of oxygen from a compound. However, the term reduction has been broadened these days to include removal of oxygen/electronegative element from a substance or addition of hydrogen/ electropositive element to a substance. According to the definition given above, the following are the examples of reduction processes:

$$2HgOS \xrightarrow{\triangle} 2\ Hg(1) + O_2(g)$$

$$2HgCl_2(aq) + SnCl_2(aq) \rightarrow Hg_2Cl_2(s) + SnCl_4(aq)$$

In reaction simultaneous oxidation of stannous chloride to stannic chloride is also occurring because of the addition of electronegative element chlorine to it. It was soon realised that oxidation and reduction always occur simultaneously (as will be apparent by re-examining all the equations given above), hence, the word "redox" was coined for this class of chemical reactions.

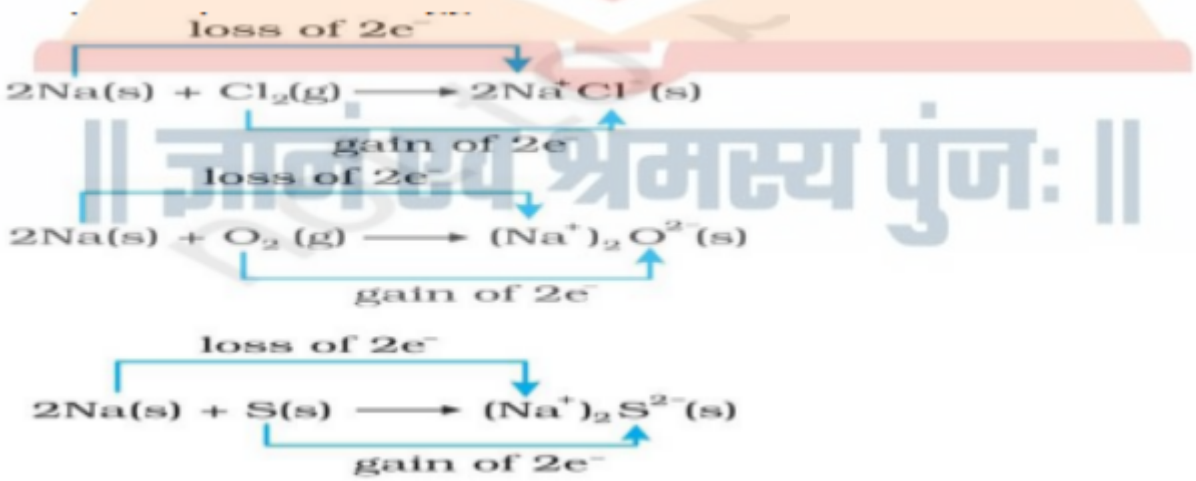
The reactions:

$$2Na(s) + Cl2g \rightarrow 2NaCl(s)$$

$$4Na(s) + O_2g \rightarrow 2Na_2o(s)$$

$$2Na(s) + S(s) \rightarrow 2Na_2S(s)$$

are redox reactions because in each of these reactions sodium is oxidised due to the addition of either oxygen or more electronegative element to sodium. Simultaneously, chlorine, oxygen and sulphur are reduced because to each of these, the electropositive element sodium has been added. From our knowledge of chemical bonding we also know that sodium chloride, sodium oxide and sodium sulphide are ionic compounds and perhaps better written as Na+ Cl–(s), $(Na^+)_2\ O^{2-}$(s), and $(Na^+)_2\ S^{2-}$(s). Development of charges on the species produced suggests us to rewrite the reactions in the following manner:



For convenience, each of the above processes can be considered as two separate steps, one involving the loss of electrons and the other the gain of electrons. As an illustration, we may further elaborate one of these, say, the formation of sodium chloride.

$$2Na(s) \rightarrow 2Na^{+}g + 2\bar{e}$$

$$Cl_2g + 2\bar{e} \rightarrow 2C\bar{I}\ (g)$$

Each of the above steps is called a half reaction, which explicitly shows involvement of electrons. Sum of the half reactions gives the overall reaction

$$2Na(s) + Cl_2(g) \rightarrow 2Na^{+}CI(s) \text{ or } 2NaCI(s)$$

Above Reactions suggest that half reactions that involve loss of electrons are called oxidation reactions. Similarly, the half reactions that involve gain of electrons are called reduction reactions. To summarise, we may mention that

Oxidation: Loss of electron(s) by any species.

Reduction: Gain of electron(s) by any species.

Oxidising agent: Acceptor of electron(s).

Reducing agent: Donor of electron(s).

i. Addition of electronegative element to a substance is known as..
   a. Oxidation
   b. Reduction
   c. Redox reaction
   d. All the above

ii. Removal of electronegative element to a substance is known as ..
   a. Oxidation
   b. Reduction
   c. Redox reaction
   d. All the above

iii. Acceptor of electrons is ...
   a. Reducing Agent
   b. Catalytic Agent
   c. Oxidising Agent
   d. None of above

iv. Donor of electrons is...
   a. Organic Agent
   b. Catalytic Agent
   c. Oxidising Agent
   d. Reducing Agent

v. Oxidation and Reduction occurs simultaneously is known as ...
   a. Exothermic reaction
   b. Endothermic reaction
   c. Redox reaction
   d. Neutralization reaction

SECTION E

*** Given Section consists of questions of 5 marks each.** **[10]**

1. Justify giving reactions that among halogens, fluorine is the best oxidant and among hydrohalic compounds, hydroiodic acid is the best reductant.

2.

While sulphur dioxide and hydrogen peroxide can act as oxidising as well as reducing agents in their reactions, ozone and nitric acid act only as oxidants. Why?

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