Redox Chemistry Questions with Solutions

Q1. For the redox reaction, $MnO_4^- + C_2O_4^{2-} + H^+ \rightarrow Mn^{2+} + CO_2 + H_2O$
The correct coefficients of the reactants for the balanced equation are-

a) 16, 5, 2
b) 2, 5, 16
c) 2, 16, 5
d) 5, 16, 2

Correct Answer: (b) 2, 5, 16

Q2. Which oxidation states of phosphorus can be found in the disproportionation reaction below?
$P_2H_4 \rightarrow PH_3 + P_4H_2$

a) P_4^{2+}
b) P^{3+}
c) P_2^{4-}
d) P^{2-}

Correct Answer: (c)

Q3. What is the oxidation number of alkali metals in its compounds?

Answer. +1.

Q4. What is the fundamental principle of balancing redox reactions using the ion-electron method?

Answer. The number of electrons lost during oxidation is equal to the number of electrons gained during reduction.

Q5. Would you use an oxidizing agent or reducing agent in order for the following reactions to occur?

a) $ClO_3^- \rightarrow ClO_2$
b) $SO_4^{2-} \rightarrow S^{2-}$
c) $Mn^{2+} \rightarrow MnO_2$
d) $Zn \rightarrow ZnCl_2$

Answer.

a) $ClO_3^- \rightarrow ClO_2$ reducing agent

b) $SO_4^{2-} \rightarrow S^{2-}$ reducing agent
c) $Mn^{2+} \rightarrow MnO_2$ oxidizing agent
d) $Zn \rightarrow ZnCl_2$ oxidizing agent

Q6. How would you know whether a redox reaction is taking place in an acidic/alkaline or neutral medium?

Answer. When H^+ or any acid appears on either side of a chemical equation, the reactions occur in an acidic medium. The solution is basic if OH^- or any base appears on either side of a chemical equation. If there is no H^+, OH^-, acid, or base in the chemical equation, the solution is neutral.

Q7. Justify the reaction:
$2Cu_2O(S) + Cu_2S\ (s) \rightarrow 6Cu\ (s) + SO_2\ (g)$ is a redox reaction. Identify the species oxidized, and reduced, which acts as an oxidant and which acts as a reductant.

Answer.

$$\overset{+1\ \ -2}{2Cu_2O(S)} + \overset{+1\ \ -2}{Cu_2S\ (s)} \rightarrow \overset{0}{6Cu\ (s)} + \overset{+4\ -2}{SO_2}$$

Copper is reduced from + 1 oxidation state to 0 oxidation state in this reaction, and sulphur is oxidised from – 2 to +4 state. As a result, the reaction is a redox reaction, additionally, Cu_2O aids sulphur in Cu_2S and Cu_2O in decreasing its oxidation number. Thus, sulphur in Cu_2S is a reducing agent.

Q8. Can we store copper sulphate solution in a gold vessel?
Given $E°_{Cu2+|\ Cu}$= +0.34 V and $E°_{Au3+|\ Au}$ = +1.50V

Answer. If the following redox reactions occur, we can store CuSO4 solution in a gold vessel.
$2Au + 3Cu^{2+} \rightarrow 2Au^{3+} + 3Cu$
The cell corresponding to the preceding redox reaction can be represented as follows:
$Au\ |\ Au^{3+}\ ||\ Cu^{2+}|Cu$
Since the reduction potential of $Au^{3+}|Au$ is higher than $Cu^{2+}|Cu$. So, the copper ion will not be reduced. Therefore, copper sulphate solution can be stored in a gold vessel. Thus, the $CuSO_4$ solution can be stored in a gold vessel.

Q9. What are the highest and lowest oxidation numbers of N?

Answer. N has a maximum oxidation number of +5 because it has five electrons in the valence shell ($2s^22p^3$) and a minimum oxidation number of -3 because it can accept three more electrons to achieve the nearest inert gas (Ne) configuration.

Q10. Consider the following galvanic cell.
$Cd\ |\ Cd^{2+}\ (1M)\ ||\ H^+(1M)\ |\ H_2\ (g/atm)$
(i) Write the overall cell reaction

(ii) What do the double vertical lines denote?

Answer:
(i) The anodic reaction is
$Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$
The Cathodic reaction is
$2H^+(aq) + 2e^- \rightarrow H_2(g)$
The overall reaction is
$Cd(s) + 2H^+(aq) \rightarrow Cd^{2+}(aq) + H_2(g)$

(ii) The salt bridge that connects the oxidation and reduction half cells is represented by the double vertical lines.

Q11. Write the following redox reaction in the oxidation & reduction half-reaction.
a) $2K(s) + Cl_2(g) \rightarrow 2KCl(s)$
b) $2Al(s) + 3Cu^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Cu(s)$

Answer:
a) $2K(s) + Cl_2(g) \rightarrow 2KCl(s)$
$K(s) \rightarrow K^+(aq) + e^-$ (Oxidation)
$Cl_2(g) + 2e^- \rightarrow 2Cl^-$ (Reduction)

b) $2Al(s) + 3Cu^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Cu(s)$
$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$ (Oxidation)
$Cu^{2+} + 2e^- \rightarrow Cu(s)$ (Reduction)

Q12. HNO_3 acts only as an oxidant whereas HNO_2 acts both as an oxidant and reductant. Why?

Answer.
The oxidation number of N in HNO_3 = +5
The oxidation number of N in HNO_2 = +3
The maximum oxidation number that N can show is = + 5
($\because$ It has only 5 valance electrons $2s^2 2p^3$)
The Oxidation number of N in HNO_3 is maximum and it can only decrease. Therefore HNO_3 can act only as an oxidant. The minimum Oxidation number of N is -3.
Thus HNO_2 in which the oxidation number of N is +3 can decrease as well as increase. Thus HNO_2 can act as an oxidant as well as a reductant.

Q13. Identify the species being oxidized and reduced in each of the following reactions:
a) $Cr^+ + Sn^{4+} \rightarrow Cr^{3+} + Sn^{2+}$
b) $3Hg^{2+} + Fe(s) \rightarrow 3Hg_2 + 2Fe^{3+}$
c) $2As(s) + 3Cl_2(g) \rightarrow 2AsCl_3$

Answer.

a) $Cr^{+} + Sn^{4+} \rightarrow Cr^{3+} + Sn^{2+}$
   Cr^{+}: oxidized, Sn^{4+}: reduced

b) $3Hg^{2+} + Fe\ (s) \rightarrow 3Hg_2 + 2Fe^{3+}$
   Hg^{2+}: reduced, Fe: oxidized

c) $2As\ (s) + 3Cl_2\ (g) \rightarrow 2AsCl_3$
   As: oxidized, Cl_2: reduced

Q14. Calculate the oxidation number of underlined elements in the followings:

a) $\mathbf{Na_2\underline{B}_4O_7}$
b) $\mathbf{H_4\underline{P}_2O_7}$
c) $\mathbf{Ca\underline{O}_2}$
d) $\mathbf{Na\underline{B}H_4}$
e) $\mathbf{H_2\underline{S}_2O_7}$

Answer:

a) $Na_2\underline{B}_4O_7$
   Let's assume the oxidation number of B is x.
   Oxidation number of H = +1
   Oxidation number of O = -2
   Then we have: 2 + 4(x) – 14 = 0
   ⇒4x = 12
   ⇒x = 3
   Hence, the oxidation number of B is +3.

b) $H_4\underline{P}_2O_7$
   Let's assume the oxidation number of P is x.
   Oxidation number of H = +1
   Oxidation number of O = -2
   Then we have: 4(+1) + 2(x) + 7 (-2) = 0
   ⇒ 4 + 2x - 14 = 0
   ⇒ 2x - 10 = 0
   ⇒ 2x = +10
   ⇒ x = +5
   Hence, the Oxidation number of P is +5.

c) $Ca\underline{O}_2$
   Let's assume the oxidation number of O is x.
   The oxidation number of Ca = +2
   Then we have: 1(+2) + 2(x) = 0
   ⇒ 2 + 2x = 0

$\Rightarrow 2x = -2$
$\Rightarrow x = -1$
Hence, the Oxidation number of O is -1

d) $Na\underline{B}H_4$
Let's assume the oxidation number of B is x.
Oxidation number of Na = +1
Oxidation number of H = -1
Then we have: $1(+1) + 1(x) + 4(-1) = 0$
$\Rightarrow 1 + x - 4 = 0$
$\Rightarrow x - 3 = 0$
$\Rightarrow x = +3$
Hence, the Oxidation number of B is +3.

e) $H_2\underline{S_2}O_7$
Let's assume the oxidation number of S is x.
Oxidation number of O = -2
Oxidation number of H = +1
Then we have: $2(+1) + 2(x) + 7(-2) = 0$
$\Rightarrow 2 + 2x - 14 = 0$
$\Rightarrow 2x - 12 = 0$
$\Rightarrow x = +6$
Hence, the Oxidation number of S is +6.

Q15. Balance the following equation by the ion-electron method
$Zn\ (s) + NO_3^- \rightarrow Zn^{2+}\ (aq) + NH_4\ (aq) + H_2O$ (In acid solution)

Answer. Oxidation half-reaction-
$Zn \rightarrow Zn^{2+}\ (aq)$
$Zn\ (s) \rightarrow Zn^{2+} + 2e^-$ (To balance charge) [equation 1]

Reduction half-reaction
$NO_3^-\ (aq) \rightarrow NH_4^+\ (aq)$
$NO_3^-\ (aq) \rightarrow NH_4^+ + 3H_2O\ (l)$ (To balance O atom)
$NO_3^-\ (aq) + 10H^+\ (aq) \rightarrow NH_4^+(aq) + 3H_2O\ (l)$ (To balance H atom)
$NO_3^-\ (aq) + 10H^+\ (aq) + 8e^- \rightarrow NH_4^+(aq) + 3H_2O\ (l)$
Multiply eq. 1 by 4 to equalise the number of electrons in both the half reactions-
$4Zn\ (s) \rightarrow 4Zn^{2+} + 8e^-$
Add both the half reaction-
$4Zn\ (s) + NO_3^-\ (aq) + 10H^+\ (aq) \rightarrow 4Zn^{2+} + NH_4^+(aq) + 3H_2O\ (l)$

Practise Questions on Redox

Q1. Determine the oxidation number of the elements in each of the following compounds:

a) H_2CO_3
b) N_2
c) $Zn(OH)_4^{2-}$
d) NO_2^-
e) LiH

Answer. The oxidation number of the elements in each of the following compounds are as follows-

a) H_2CO_3
   H: +1, O: -2, C: +4
b) N_2
   N: 0
c) $Zn(OH)_4^{2-}$
   Zn: 2+, H: +1, O: -2
d) NO_2^-
   N: +3, O: -2
e) LiH
   Li: +1, H: -1

Q2. Write the balanced half-reactions of the following reactions:

a) $NiO_2 + 2H_2O + Fe \rightarrow Ni(OH)_2 + Fe(OH)_2$ in basic solution
b) $CO_2 + 2N_2HOH \rightarrow CO + N_2 + 3H_2O$ in basic solution

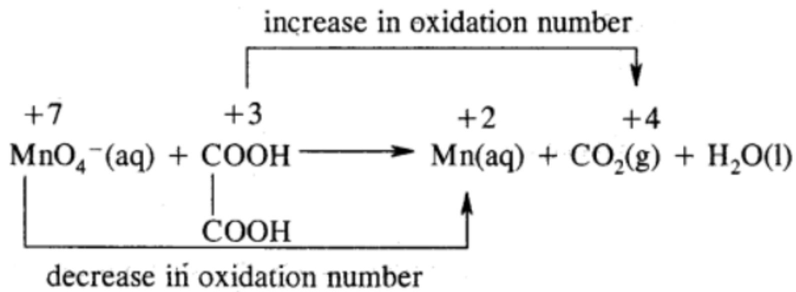
Answer. The balanced half-reactions of the following reactions are as follows-

a) $NiO_2 + 2H_2O + Fe \rightarrow Ni(OH)_2 + Fe(OH)_2$ in basic solution
   $2H_2O + NiO_2 + 2e^- \rightarrow Ni(OH)_2 + 2OH^-$
   $2OH^- + Fe \rightarrow Fe(OH)_2 + 2e^-$

b) $CO_2 + 2N_2HOH \rightarrow CO + N_2 + 3H_2O$ in basic solution
   $CO_2 + H_2O + 2e^- \rightarrow CO + 2OH^-$
   $2OH^- + 2NH_2OH \rightarrow N2 + 2e^- + 4H_2O$

Q3. Balance the following equation in an acidic medium by the oxidation number method.
$MnO_4^- + (COOH)_2 \rightarrow Mn^{2+} (aq) + CO_2 + H_2O$

Answer.



$2MnO_4^- + 5(COOH)_2 \rightarrow 2Mn^{2+}\,(aq) + 10CO_2 + 8H_2O$
$2MnO_4^- + 5(COOH)_2 + 6H^+ \rightarrow 2Mn^{2+}\,(aq) + 10CO_2 + 8H_2O.$

Q4. How many grams of $K_2Cr_2O_7$ is required to oxidize Fe^{2+} present in 15.2 gm of $FeSO_4$ to Fe^{3+} if the reaction is carried out in an acidic medium.

Answer:
The balanced chemical equation for the redox reaction is
$K_2Cr_2O_7 + 6FeSO_4 + 7H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + 7H_2O$

The balanced equation shows that 6 moles of $FeSO_4$ = 1

$K_2Cr_2O_7$ oxidises a mole of $K_2Cr_2O_7$ or 6 × 152 gm of $FeSO_4$ = 294 gm
or
15.2 gm of $FeSO_4$ oxidised by $K_2Cr_2O_7$

Q5. How many millimoles of potassium dichromate are required to oxidize 24 cm^3 of 0.5 M mohr's salt solution in an acidic medium.

Answer.
No. of millimoles of $K_2Cr_2O_7$ present in 24 cm^3 of 0.5 m solution = 24 × 0.5 = 12
The balanced chemical equation for the redox reaction is
$K_2Cr_2O_7 + 6(NH_4)_2SO_4.FeSO4.6H_2O + 7H_2SO_4 \rightarrow K_2SO_4 + 6(NH_4)_2SO_4 + 3Fe_2(SO_4)_3 + Cr_2(SO_4)_3 + 43H_2O$

From the balanced equation, 6 moles of mohr's salt is oxidised by $K_2Cr_2O_7$ = 1 mole
∴ 12 millimoles of mohr's salt will be oxidised by
$K_2Cr_2O_7$ = 1/6 × 12 = 2 millimoles.