Chemistry Worksheets Class 11 on Chapter 8 Redox Reactions with Answers - Set 2

Q-1: In the subsequent reaction, $K + O_2 \rightarrow KO_2$, the oxidation state of oxygen is altered from __________.

a) 0 to -1
b) 0 to -2
c) 0 to -½
d) 0 to +1

Answer: c) 0 to -½
Explanation: A superoxide, KO_2, has an oxidation state of -½, while O_2 has an oxidation state of zero. As a result, the oxidation state shifts from 0 to -½.

Q-2: Which substance never exhibits a negative oxidation number?
a) O
b) Fe
c) S
d) F
Answer: b) Fe

Q-3: Which statement regarding the salt bridge is accurate?
a) It completes the circuit
b) It maintains neutrality
c) It contains salt
d) All of the above

Answer: d) All of the above
Explanation: A salt bridge is a U-tube that contains inert electrolyte solutions, such as KCl or KNO_3 (salts), in agar-agar or gelatin.
Here are a few of the salt bridge's functions:
i) To complete the circuit.
ii) To keep the solutions of the two half-cells electrically neutral.
iii) To stop the solutions from the two half-cells from mixing.
The above-mentioned functions state that option (d) is the correct answer.

Q-4: An electrochemical cell has two half cell reactions as,
$B^{2+} + 2e^- \rightarrow B$; $E^0 = 0.34$ V
$Y \rightarrow Y^{2+} + 2e^-$; $E^0 = 2.37$ V

The cell voltage will be
a) 2.71 V
b) 2.03 V
c) -2.71 V
d) -2.03 V

Answer: a) 2.71 V
Explanation: By combining the oxidation and reduction potentials of the two electrodes, the cell voltage can be obtained:
$E^o_{cell} = E^o_{oxidation} + E^o_{reduction}$
$= 2.37\ V + 0.34\ V$
$= 2.71\ V$

Q-5: Which of the following is the most reactive metal?
a) K
b) Zn
c) Ni
d) Ag

Answer: a) K
Explanation: According to the reactivity series, potassium is the most reactive among all the metals given.

Q-6: How does sulphur show a disproportionation reaction? Give an example.
Answer: Sulphur exhibits the disproportionation reaction in the manner described below:
$S_8 + 12OH^-(aq) \rightarrow 4S^{2-} + 2S_2O_3^{2-} + 6H_2O$
We can deduce from the reaction that sulphur is in a zero oxidation state in S_8 and changes to a -2 in S^{2-} and a +2 oxidation state in $S_2O_3^{2-}$. Sulphur is undergoing disproportionation because it is both being reduced and oxidised.

Q-7: Chloride and chlorate ions are produced when chlorine gas is passed through a concentrated alkali solution. Identify the reaction and obtain a balanced chemical equation for it.
Answer: The balanced chemical reaction is shown below:
$Cl_2 + 2OH^-(aq) \rightarrow ClO^- + Cl^- + H_2O$

This reaction is a disproportionation reaction because chlorine undergoes an oxidation state change from zero in Cl_2 to -1 in Cl^- and +1 in ClO^-. When an atom of the same element undergoes both reduction and oxidation simultaneously, this is known as a disproportionation reaction.

Q-8: Give a classical definition of oxidation along with some suitable examples.
Answer: According to the classical concept, oxidation involves the following:

i) Addition of oxygen or
ii) Addition of electronegative element or
iii) Removal of hydrogen or
iv) Removal of electropositive element

For example,
Addition of oxygen: $2Mg + O_2 \rightarrow 2MgO$
Addition of electronegative element: $Mg + Cl_2 \rightarrow MgCl_2$
Removal of hydrogen: $MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$
Removal of electropositive element: $2KI + Cl_2 \rightarrow 2KCl + I_2$

In conclusion, oxidation is the process of adding oxygen or another electronegative element to a substance or removing hydrogen or another electropositive element from a substance.

Q-9: Write the anodic and cathodic reactions for the galvanic cell after taking into account the following reactions.
i) $2Fe^{3+}(aq) + 2Cl^-(aq) \rightarrow 2Fe^{2+}(aq) + Cl_2(g)$
ii) $2Cr(s) + 3Cu^{2+} \rightarrow 3Cu(s) + 2Cr^{3+}$

Answer: Oxidation reaction takes place at the anode and reduction reaction takes place at the cathode.
Below are the anodic and cathodic reactions for the given reactions:
i) At anode: $2Cl^- \rightarrow Cl_2 + + 2e^-$
At cathode: $2Fe^{3+} + 2e^- \rightarrow 2Fe^{2+}$
ii) At anode: $2Cr \rightarrow 2Cr^{3+} + 6e^-$
At cathode: $3Cu^{2+} + 6e^- \rightarrow 3Cu$

Q-10: The reduction potentials of the following reactions are:
$Cl_2 + 2e^- \rightarrow 2Cl^-$; $E^o = 1.36$ V
$F_2 + 2e^- \rightarrow 2F^-$; $E^o = 2.87$ V
Which is a better oxidising agent?

Answer: F_2 is a better oxidising agent.
Explanation: An oxidising agent reduces itself but causes the oxidation of other species. Because reduction potential is the tendency of a substance to get reduced itself, the greater the reduction potential, the better the oxidising agent.
F_2 is a better oxidising agent than Cl_2 because it has more reduction potential.

Q-11: Show that the following process is a redox one.
$3CuO + 2NH_3 \rightarrow 3Cu + N_2 + 3H_2O$

Answer: The Cu in CuO is in the +2 oxidation state in the reaction described above, and it is changing to the oxidation state of 0 in Cu, indicating that it is getting reduced. Similarly, the N in NH_3 is in the -3 oxidation state, changing to the 0 oxidation state in N_2, indicating that it is being oxidised. This is a redox change as simultaneous oxidation and reduction are taking place.

Q-12: What is the element's oxidation number in each of the following species?

a) K<u>I</u>$_3$

b) <u>C</u>O_3^{2-}

c) <u>N</u>H_3

d) H<u>Cl</u>O

e) H<u>P</u>O_4^{2-}

Answer: Let's first assume that each atom has an oxidation number of x to calculate the oxidation number of each one that is underlined. The oxidation number method can now be used to calculate the oxidation number.

a) K<u>I</u>$_3$: +1+3(x) = 0

On solving, x = -⅓

Hence, the oxidation number of I in KI_3 is -⅓.

b) <u>C</u>O_3^{2-}: x +3(-2) = -2

On solving, x= +4

Hence, the oxidation number of C in CO_3^{2-} is +4.

c) <u>N</u>H_3: x+ 3(1) = 0

On solving. x = -3

Hence, the oxidation number of N in NH_3 is -3.

d) H<u>Cl</u>O: +1+x-2 = 0

On solving, x = +1

Hence, the oxidation number of Cl in HClO is +1.

e) H<u>P</u>O_4^{2-}: +1+x+4(-2)= -2

On solving, x = +5

Hence, the oxidation number of P in HPO_4^{2-} is +5.

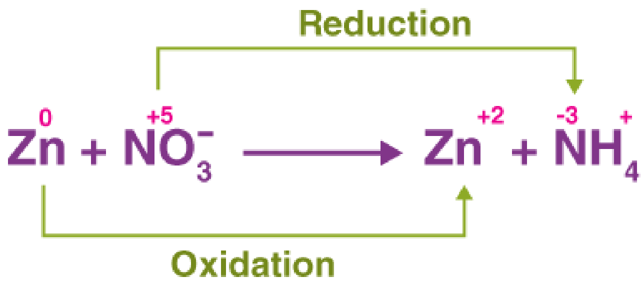
Q-13: Give an example of a compound in which the oxidation number is fractional.

Answer: Carbon Suboxide (C_3O_2) is a compound in which the oxidation number is fractional. In C_3O_2, the oxidation number of two 'C' atoms in the terminal position is +2, while that of the atom in the centre is zero. The average oxidation number of carbon atoms is (2+2+0)/3 = 4/3.

Q-14: Balance the equation by ion electron method.

$Zn + NO_3^- \rightarrow Zn^{2+} + NH_4^+$ (in basic medium)

Answer:



i)
ii) $Zn \rightarrow Zn^{2+}$ (Oxidation half-reaction)
$NO_3^- \rightarrow NH_4^+$ (Reduction half-reaction)

iii) Balancing e^-
$Zn \rightarrow Zn^{2+} + 2e^-$
$NO_3^- + 8e^- \rightarrow NH_4^+$

iv) Balancing 'O'
$Zn \rightarrow Zn^{2+} + 2e^-$
$NO_3^- + 8e^- \rightarrow NH_4^+ + 3H_2O$

v) Balancing 'H'
$NO_3^- + 8e^- + 10H_2O \rightarrow NH_4^+ + 3H_2O + 10OH^-$
$Zn \rightarrow Zn^{2+} + 2e^-$
$NO_3^- + 7H_2O + 8e^- \rightarrow NH_4^+ + 10OH^-$
Now,
$[Zn \rightarrow Zn^{2+} + 2e^-] \times 4$
$NO_3^- + 7H_2O + 8e^- \rightarrow NH_4^+ + 10OH^-$

$4Zn + NO_3^- + 7H_2O \rightarrow 4Zn^{2+} + NH_4^+ + 10OH^-$

Q-15: What do you mean by non-metal displacement reactions? Explain with an example.
Answer: The non-metal displacement redox reactions include an infrequent reaction involving oxygen displacement and hydrogen displacement.

Examples include the displacement of hydrogen from cold water by all alkali metals and some alkaline earth metals (Ca, Sr, and Ba), which are excellent reductants. Below is the reaction of an alkali metal, Na, with cold water:

$2Na + H_2O \rightarrow 2NaOH + H_2$

Q-16: Explain the combination reaction.
Answer: The chemical reactions in which two or more substances combine to form a single substance are called combination reactions.
A combination reaction may be expressed as $A + B \rightarrow AB$
An example of such a reaction is $C + O_2 \rightarrow CO_2$

Q-17: What is the electrochemical series? Write two applications for it.
Answer: The term "electrochemical series," also known as "activity series," refers to a list that describes the arrangement of elements in ascending order of their electrode potential values. By comparing the potential of various electrodes to a standard hydrogen electrode, a series has been established.
The two applications of the electrochemical series are:
a) Activity of metals: Lower the standard reduction potential value, the higher will be the activity of the metals, as it will have a higher tendency to lose electrons.
b) Predicting whether a metal can liberate hydrogen from acid or not: All those metals whose reduction potential are less than hydrogen will have a higher tendency to lose electron than hydrogen, and so H^+ ions will gain electrons to form H_2.

Q-18: Write formulas for the following compounds.
a) Mercury (II) chloride
b) Tin (IV) oxide
c) Iron (III) sulphate
d) Chromium (III) oxide

Answer:
a) $HgCl_2$
b) SnO_2
c) $Fe_2(SO_4)_3$
d) Cr_2O_3

Q-19: Arrange the following metals in the order in which they displace each other from the solutions of their salt.
Ca, Pb, Sn, H, Fe, Ni

Answer: The less reactive metal can be displaced from its salt solution by the more reactive metal.
Metal reactivity is arranged as follows: Ca > Fe > Ni > Sn > Pb > H
Because calcium is more reactive than all other metals and can thus displace them from their salt solutions, this is according to the order of reactivity.

This results in the following metals being displaced from their salt solution in the following order: Ca > Fe > Ni > Sn > Pb > H.

Q-20: Predict whether the reaction between the following is feasible**.**

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

b) $Br_2(aq) + Fe^{2+}(aq)$

c) $Fe^{3+}(aq) + I^-(aq)$

Answer:

a) The reaction occurs in the following way:

$Ag(s) + 2Fe^{3+}(aq) \rightarrow Ag^+(aq) + 2Fe^{2+}(aq)$

Oxidation half-reaction: $Ag(s) \rightarrow Ag^+ + e^-$ $E^o = -0.80$ V

Reduction half-reaction: $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}$ (aq) $E^o = 0.77$ V

Overall reaction: $Ag(s) + 2Fe^{3+}(aq) \rightarrow Ag^+(aq) + 2Fe^{2+}(aq)$

$E^o_{cell} = -0.80 \text{ V} + 0.77 \text{ V} = -0.03 \text{ V}$

Since E^o_{cell} is negative, the reaction is not feasible.

b) Oxidation half-reaction: $Fe^{2+}(aq) \rightarrow Fe^{3+} + e^-$ (aq) $E^o = -0.77$ V

Reduction half-reaction: $Br_2(aq) + 2e^- \rightarrow 2Br^-$ (aq) $E^o = 1.09$ V

Overall reaction: $Br_2(aq) + 2Fe^{2+}(aq) \rightarrow 2Br^-(aq) + 2Fe^{3+}(aq)$

$E^o_{cell} = -0.77 \text{ V} + 1.09 \text{ V} = 0.32 \text{ V}$

Since E^o_{cell} is positive, the reaction is feasible.

c) Oxidation half-reaction: $2I^-(aq) \rightarrow I_2(s) + 2e^-$ (aq) $E^o = -0.54$ V

Reduction half-reaction: $Fe^{3+}(aq) \rightarrow Fe^{2+}(aq) + e^-(aq)$ $E^o = 0.77$ V

Overall reaction: $2I^-(aq) + 2Fe^{3+}(aq) \rightarrow I_2(s) + 2Fe^{2+}(aq)$

$E^o_{cell} = -0.54 + 0.77 \text{ V} = 0.23 \text{ V}$

Since the E^o_{cell} is positive, the reaction is feasible.