Redox Reactions

7
Chapter

1. **Assign oxidation numbers to the underlined elements in each of the following species**:

(a) NaH_2PO_4

Ans: P's oxidation number will be x.

We are aware of this.

Oxidation number of $Na = +1$

Oxidation number of $H = +1$

Oxidation number of $O = -2$

$(+1)(+1)(x)(-2) \quad NaH_2PO_4$

Then there's

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

$1+2+x-8=0$

$x=+5$

As a result, P's oxidation number is $+5$

(b) $NaHSO_4$

$(+1)(+1)(x)(-2)$

Ans: $NaHSO_4$

Then there's

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

$1+1+x-8=0$

$x=+6$

As a result, S's oxidation number is $+6$

(c) $H_4P_2O_7$

Ans: $H_4P_2O_7$

Then, there's

$4(+1)+2(x)+7(-2)=0$

$4+2x-14=0$

$2x=+10$

$x=+5$

As a result, P's oxidation number is $+5$

(d) K_2MnO_4

Ans: Then, there's

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

$2+x-8=0$

$x=+6$

As a result, Mn 's oxidation number is $+6$

(e) CaO_2

Ans:

Then, there's

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

$2+2x=0$

x=-1

As a result, O's oxidation number is -1

(f) $NaBH_4$

Ans: Then, there's

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

$1+x-4=0$

$x=+3$

As a result, B's oxidation number is $+3$

(g) $H_2S_2O_7$

Ans: $H_2S_2O_7$

Then, there's

$2(+1)+2(x)+7(-2)=0$

$2+2x-14=0$

$2x=12$

$x=+6$

As a result, S's oxidation number is $+6$

(h) $KAl(SO_4)_2.12H2O$

Ans: $KAl(SO_4)_2.12H_2O$

Then, there's

$1(+1)+1(+3)+2(x)+8(-2)+24(+1)+12(+2)=0$

$1+3+2x-16+24-24=0$

$2x=12$

$x = +6$

As a result, S's oxidation number is $+6$

Because water is a neutral molecule, we can disregard it. The sum of all atoms in the water molecule's oxidation numbers can then be considered as zero. As a result of disregarding the water molecule, we now have

2. What are the oxidation numbers of the underlined elements in each of the following and how do you rationalize your results?

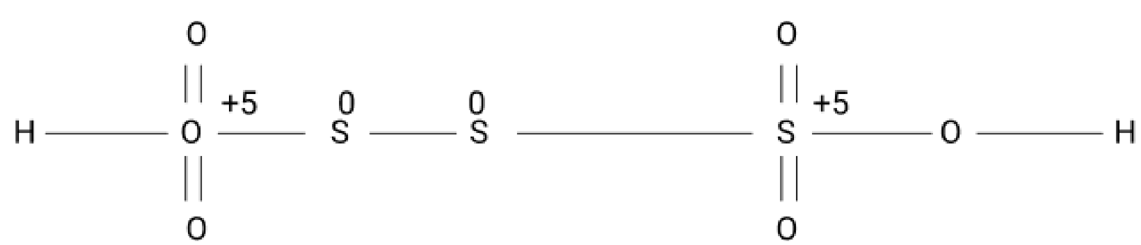
(a) KI_3

Ans: In KI_3 K has an oxidation number (O.N.) of one. As a result, I's average oxidation number is $\frac{1}{3}.0.N$, on the other hand, cannot be fractional. To determine the oxidation states, we must first study the structure of KI_3.

An iodine atom makes a coordinate covalent link with an iodine molecule in a KI_3 molecule.

$$\overset{+1}{K^+}\left[\overset{0}{I}-\overset{0}{I}\leftarrow\overset{-1}{I}\right]$$

As a result, the O.N. of the two atoms that make up the I_2 molecule in a KI_3 molecule is 0, whereas the O.N of the I atom that makes up the coordinate bond is -1.



Two of the four S atoms have an O.N. of $+5$, whereas the other two have an O.N. of 0

(b) $H_2S_4O_6$

Ans: Now, $2(+1)+4(x)+6(-2)=0$

$\Rightarrow 2+4x-12=0$

$\Rightarrow 4x=10$

However, O.N. cannot be fractional. Hence, S must be present in different oxidation states in the molecule.

The O.N. of two of the four atoms is and the O.N. of the other two atoms is 0.

(c) Fe_3O_4

Ans: When the $O . N$. of O is set to -2, the O.N. of Fe is found to be $+2\frac{2}{3}$. O.N., on the other hand, cannot be fractional.

One of the three Fe atoms in this example has an O.N. of $+2$, whereas the other two Fe atoms have an O.N. of $+3$.

$FeO^{+2}Fe_2O_3$

(d) CH_3CH_2OH

Ans: $x+1-2$

C_2H_6O

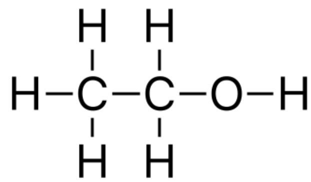
$2(x)+6(+1)+1(-2)=0$

$2x+6-2=0$

$x=-2$

This molecule's two carbon atoms are found in two separate settings. As a result, their oxidation numbers cannot be the same.

As a result, C has the oxidation states of -3 and -1.



(e) CH_3COOH

Ans:

$x \quad +1 \quad -2$

$C_2H_4O_2$

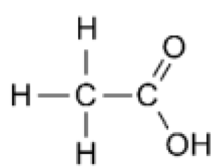
$2(x)+4(+1)+2(-2)=0$

$2x+4-4=0$

$x=0$

The average O.N. of C, on the other hand, is 0 . This molecule's two carbon atoms are found in two separate settings. As a result, their oxidation numbers cannot be the same.

In CH_3COOH, C has the oxidation states of $+3$ and -3.



Acetic acid

3. **Justify that the following reactions are redox reactions:**

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

Ans: Let's write the oxidation number of each element in the reaction as follows:

$$CuO(s)+H_2(g) \rightarrow Cu(s)\ +1\ \ 2-$$

Cu 's oxidation number falls from $+2$ in CuO to 0 in Cu, implying that CuO is reduced to Cu. In addition, the oxidation number of H in H_2 increases from 0 to $+1$ in H_2O, indicating that H_2 is oxidized to H_2O. As a result, this is a redox reaction.

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

Ans: Let's write the oxidation number of each element in the reaction as follows:

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

Fe 's oxidation number falls from $+3$ in Fe_2O_3 to 0 in Fe, implying that Fe_2O_3 is reduced to Fe. The oxidation number of C, on the other hand, increases from $+2$ in CO to $+4$ in CO_2, indicating that CO is oxidized to CO_2. As a result, the reaction in question is a redox reaction.

(c) $4BCl_3(g)+3LiAlH_4(s) \rightarrow 2B_2H_6(g)\ +3LiCl(s)\ \ 3+AlCl_3(s)$

Ans: Let's write the oxidation number of each element in the reaction as follows:

$$4BCl_3(g)+3LiAlH_4(s) \rightarrow 2B_2H_6(g)\ +3LiCl(s)\ \ 3+AlCl_3(s)$$

The oxidation number of B drops from $+3$ in BCl_3 to -3 in B_2H_6 in this reaction. BCl_3 is reduced to B_2H_6 in this way. In addition, the oxidation number of H in $LiAlH_4$ increases to -1 in B_2H_6, indicating that LiAlH 4 is oxidized to B_2H_6. As a result, the reaction in question is a redox reaction.

(d) $2K(s)+F_2(g) \rightarrow 2K\ +F\ \ (s)$

Ans: Let's write the oxidation number of each element in the reaction as follows:

$$2K(s)+F_2(g) \rightarrow 2K\ +F\ \ (s)$$

The oxidation number of K increases from 0 in to $+1$ in $K F$ in this reaction, indicating that K is oxidized to KF. The oxidation number of F, on the other hand, decreases from 0 in F_2 to -1 in KF, indicating that F_2 is reduced to KF.

As a result, the preceding reaction is a redox reaction.

(e) $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$

Ans:

Let's write the oxidation number of each element in the reaction as follows:

$-3 \quad +1 \quad 0$

$4NH_3(g) + 5O_2(g) \rightarrow 4 \ NO(g) + 6H_2O(g)$

The oxidation number of N rises from -3 in NH_3 to $+2$ in NO in this case. The oxidation number of O_2 drops from 0 in O_2 to -2 in NO and H_2O, indicating that O_2 is reduced.

As a result, the reaction in question is a redox reaction.

4. Fluorine reacts with ice and results in the change :

$H_2O(s) + F_2(g) \rightarrow HF(g) + HOF(g)$

Justify that this reaction is a redox reaction.

Ans:

Let's write the oxidation number of each atom in the reaction above its symbol as follows:

$+1 \quad -2 \quad 0 \quad 1 +1 \cdot 1 \ 2 +1$

$H_2O(s) + F_2(g) \rightarrow HF(g) + HOF(g)$

Here, we have observed that the oxidation number of F increases from 0 in F_2 to $+1$ in HQF . Also, the oxidation number decreases from 0 in F_2 to -1 in HF. Thus, in the above reaction, F is both oxidized and reduced. Hence, the given reaction is a redox reaction.

5. **Calculate the oxidation number of Sulphur, chromium, and nitrogen in H_2SO_5, $Cr_2O_7^{2-}$ and NO_3^-. Suggest structure of these compounds. Count for the fallacy.**

Ans:

$\overset{+1}{}\quad \overset{x}{}\ \overset{-2}{}$

(i) H_2SO_5

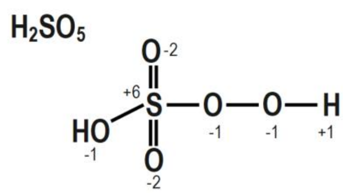
$2(+1)+1(x)+5(-2)=0$

$2+x-10=0$

$x=+8$

S's O.N., on the other hand, cannot be $+8$. S has six electrons in its valence shell. As a result, S's O.N. cannot be greater than $+6$.

The structure of H_2SO_5 is depicted in the diagram below.



$2(\mathrm{H})+1(\mathrm{S})+3(0)+2(0$ in peroxy linkage $2(+1)+1(x)+3(-2)+2(-1)=0$

$2+x-6-2=0$

$x=+6$

As a result, S's O.N. is $+6$.

(ii)

$x \quad -2$

$Cr_2O_7^{2-}$

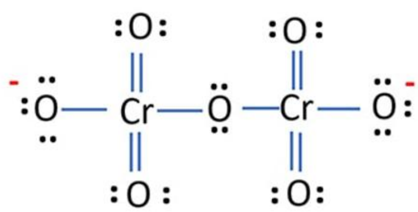
$2(x)+7(-2)=-2$

$2x-14=-2$

$x=+6$

The O.N. of Cr in $\mathrm{Cr_2O_7}^{2-}$ is not a fallacy in this case.

The structure of $\mathrm{Cr_2O_7}^{2-}$ is depicted in the diagram below



Each of the two Cr atoms here have an O.N. of $+6$.

(iii) $\mathrm{NO_3}$

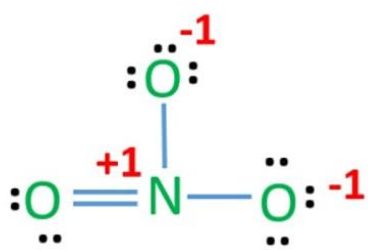
$1(x)+3(-2)=-1$

$x-6=-1$

$x=+5$

The O.N. of N in NO_3° is not a fallacy in this case.

The structure of NO_3° is depicted in the diagram below



The O.N. value of the N atom is +5.

6. Write the formulae for the following compounds:

(a) Mercury (II) chloride

Ans: H_gCl_2

(b) Nickel (II) sulphate

Ans: $NiSO_4$

Ans: SnO_2 (c) Tin (IV) oxide

(d) Thallium(I) sulphate

Ans: Tl_2SO_4

(e) Iron (III) sulphate

Ans: $Fe_2(SO_4)_3$

(f) Chromium (III) oxide

Ans: Cr_2O_3

7. Suggest a list of the substances where carbon can exhibit oxidation states from –4 to +4 and nitrogen from –3 to +5.

Ans:

Substance	O.N. of carbon
CH_2Cl_2	0
$ClC \equiv CCl$	+1
$HC = CH$	–1

$CHCl_3, CO$	+2
$CHCl_3$	–2
Cl_3C-CCl_3	+3
H_3C-CH_3	–3
CCl_4, CO_2	+4
CH_4	–4

The substances where nitrogen can exhibit oxidation states from –3 to +5 are listed in the following table.

Substance	O.N. of carbon
N_2	0
N_2O	+1
N_2H_2	–1
NO	+2
N_2H_4	–2
N_2O_3	+3
NH_3	–3

NO_2	+4
N_2O_5	+5

The oxidation number (O.N.) of S in sulphur dioxide (SO_2) is $+4$, while the O.N. of S can range from $+6$ to -2.

As a result, SO_2 can function as both an oxidising and a reducing agent.

The O.N. of O in hydrogen peroxide (H_2O_2) is -1, and the range of $O.N.$ that O can have is 0 to -2. The oxidation values $+1$ and $+2$ are also possible for O.

As a result, H_2O_2 can function as both an oxidizing and a reducing agent.

As a result, in this scenario, the O.N. of O can only drop. As a result, O_3 serves solely as an oxidant.

The O.N. of N in nitric acid (HNO_3) is $+5$, and the range of $O.N.$ that N can have is from $+5$ to -3. As a result, in this scenario, the O.N. of N can only drop. As a result, HNO, serves solely as an oxidant.

8. 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?

Ans: The S atom in SO_2 has +4 oxidation number. The minimum and maximum oxidation numbers of S are -2 and +6 respectively. Hence, in SO_2, S can increase and decrease its oxidation number. Hence, SO_2 is an oxidizing agent as well as reducing agent.

The O atom in hydrogen peroxide has oxidation number of -1. The minimum and maximum oxidation numbers of O are -2 and 0 respectively. Hence, hydrogen peroxide is oxidant as well as reluctant.

In ozone, O atom has oxidation number of 0. It can decrease its oxidation number to -1 or -2 but cannot increase it. Hence ozone is an oxidizing agent.

In nitric acid, N has oxidation number of +5 which is maximum. N can only decrease its oxidation number. Hence, nitric acid is an oxidizing agent.

9. Consider the reactions:

(a) $6CO_2(g) + 6H2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$

(b) $O_3(g) + H_2O_2(1) \rightarrow H_2O(1) + 2O_2(g)$

Why it is more appropriate to write these reactions as:

(a) $6CO_2(g) + 12H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6H_2O(1) + 6O2(g)$

Ans: The H_2 produced in step 1 reduces CO_2, thereby producing glucose $(C_6H_{12}O_6)$ and H_2O.

$$6CO_{2(g)} + 12H_{2(g)} \rightarrow C_6H_{12}O_{6(s)} + 6H_2O_{(i)}$$

Now, the net reaction of the process is given as: $\left.2H_2O_{(t)} \rightarrow 2H_{2(g)} + O_{2(g)}\right] \times 6$

$$\frac{6CO_{2(g)} + 12H_{2(g)} \rightarrow C_6H_{12}O_{6(s)} + 6H_2O_{(l)}}{6CO_{2(g)} + 12H_2O_{(l)} \rightarrow C_6H_{12}O_{6(g)} + 6H_2O_{(l)} + 6O_{2(g)}}$$

It is more appropriate to write the reaction as given above because water molecules are also produced in the process of photosynthesis. The path of this reaction can be investigated by using radioactive H_2O_{18} in place of H_2O

(b) $O_3(g) + H_2O_2(l) \rightarrow H_2O(1) + O_2(g)$

Ans: O_2 is produced from each of the two reactants O_3 and H_2O_2. For this reason, O_2 is written twice.

The given reaction involves two steps. First, O_3 decomposes to form O2 and O. In the second step, H_2O_2 reacts with the O produced in the first step, thereby producing H_2O and O_2

$$O_{3(g)} \rightarrow O_{2(g)} + O_{(g)}$$

$$\frac{H_2O_{2(l)} + O_{(g)} \rightarrow H_2O_{(l)} + O_{2(g)}}{2H_2O_{2(l)} + O_{3(g)} \rightarrow H_2O_{(l)} + O_{2(g)} + O_{2(g)}}$$

The path of this reaction can be investigated by using, $H_2O_2^{18}$ or O_3^{18}

10. The compound AgF_2 is an unstable compound. However, if formed, the compound acts as a very strong oxidizing agent. Why?

Ans:

Ag in AgF_2 has an oxidation state of $+2$. However, Ag 's oxidation state of $+2$ is unstable.

As a result, silver quickly takes an electron to create $Ag+$ whenever AgF_2 is formed. This helps to reduce Ag 's oxidation state from $+2$ to $+1$, which is a more stable condition. As a result, AgF_2 is an extremely powerful oxidizing agent.

11. Whenever a reaction between an oxidizing agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing agent is in excess and a compound of higher oxidation state is formed if the oxidizing agent is in excess. Justify this statement giving three illustrations.

Ans: When an oxidizing agent and a reducing agent react, a lower oxidation state compound is formed if the reducing agent is in excess, and a higher oxidation state compound is formed if the oxidizing agent is in excess. As an example, consider the following:

(i) Reducing and oxidising agents, respectively, are P_4 and F_2.

When an excess of P_4 is treated with F_2, PF_3 is formed, with a positive oxidation number (O.N.) for P.

However, if P_4 is treated with an excessive amount of F_2, PF_5 is formed, with a P.N. of $+5$.

(ii) O_2 is an oxidising agent, whereas K is a reducing agent.

K_2O is generated when an excess of K reacts with O_2, with the O.N. of O being -2.

$$4K(\text{ excess }) + O_2 \rightarrow 2K_2O$$

When K reacts with an excess of O_2, however, $2K_2O_2$ is produced, with the O.N. of O being -

$$4K + O_2(\text{ excess }) \rightarrow 2K_2O_2$$

While C is a reducing agent, O_2 is an oxidizing agent.

CO is created when an excess of C is burned in the presence of inadequate O_2, with the O.N. of C being $+2$

$$C(\text{ excess }) + O_2 \rightarrow CO$$

If there is an excess of O_2 in the combustion of C, CO_2 is generated, with the O.N. of C being $+4$

$$C + O_2(\text{ excess }) \rightarrow CO_2$$

12. How do you count for the following observations?

(a) Though alkaline potassium permanganate and acidic potassium permanganate both are used as oxidants, yet in the manufacture of benzoic acid from toluene we use alcoholic potassium permanganate as an oxidant. Why? Write a balanced redox equation for the reaction.

(b) When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colorless pungent smelling gas HCl, but if the mixture contains bromide then we get red vapour of bromine. Why?

Ans: (a) Alcoholic potassium permanganate is utilized as an oxidant in the production of benzoic acid from toluene for the following reasons.

(i) In a neutral medium, OH- ions are produced in the reaction itself. As a result, the cost of adding an acid or a base can be reduced.

(ii) Because both $KMnO_4$ and alcohol are polar, they are homogenous. Because they a both organic molecules, toluene and alcohol are also homogenous.

In a homogeneous medium, reactions can proceed more quickly than in a heterogeneous one. As a result, $KMnO_4$ and toluene might react more quickly in alcohol.

For the reaction in a neutral medium, the balanced redox equation is as follows:

$$C_6H_5-CH_3(l)+2MnO_4^-(\text{alcoholic}) \rightarrow C_6H_5COO^-(\text{alcoholic})$$

$$+2MnO_2(s)+H_2O(l)+OH^-(aq)$$

When concentrated H_2SO_4 is introduced to an inorganic bromide mixture, HBr is generated at first. With the formation of red bromine vapour, HBr, as a powerful reducing agent, lowers H_2SO_4 to SO_2.

$$2NaBr+2H_2SO_4 \rightarrow 2NaHSO_4+2HBr \quad 2HBr+H_2SO_4 \rightarrow Br_2+SO_2+2H_2O \text{ (red vapour)}$$

When concentrated H_2SO_4 is added to an inorganic combination with chloride, a pungent-smelling gas (HCl') is produced. Because HCl is a poor reducing agent, it cannot convert H_2SO_4 to SO_2

$$2NaCl+2H_2SO_4 \rightarrow 2NaHSO_4+2HCl$$

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

(a) $2AgBr(s) + C_6H_6O_2\ (aq) \rightarrow 2Ag(s) + 2HBr(aq) + C_6H_4O_2(aq)$

Ans:

Oxidized substance $\rightarrow C_6H_6O_2$

Reduced substance $\rightarrow AgBr$

Oxidizing agent $\rightarrow AgBr$

Reducing agent $\rightarrow C_6H_6O_2$

(b)

$HCHO(1) + 2\left[Ag\left(NH_3\right)_2\ ^{+}\right](aq) + 3OH^{\circ}(aq) \rightarrow 2Ag(s) + HCOO^{-}(aq) + 4NH_3(aq) + 2H_2O(I)$

Ans: Oxidized substance $\rightarrow$ HCHO

Reduced substance $\rightarrow \left[Ag\left(NH_3\right)_2\ ^{+}\right]$ Oxidising agent $\rightarrow \left[Ag\left(NH_3\right)_2\ ^{+}\right]$ Reducing agent $\rightarrow HCH$

(c) $HCHO(I) + 2Cu^{2+}(aq) + 5OH^{\circ}(aq) \rightarrow Cu_2O(s) + HCOO$ **(aq)** $+3H_2O(I)$

Ans: Oxidised substance $\rightarrow$ HCHO

Reduced substance $\rightarrow Cu^{2+}$

Oxidising agent $\rightarrow Cu^{2+}$

Reducing agent $\rightarrow$ HGHO

(d) $N_2H_4(I) + 2H_2O_2$ **(I)** $\rightarrow N_2(g) + 4H_2O(I)$

Ans: Oxidised substance $\rightarrow N_2H_4$

Reduced substance $\rightarrow H_2O_2$

Oxidising agent $\rightarrow H_2O_2$

Reducing agent $\rightarrow N_2H_4$

(e) $Pb(s)+PbO_2(s)+2H_2SO_4(aq)\rightarrow 2PbSO_4(s)+2H_2O(I)$

Ans: Oxidised substance $\rightarrow$ Reduced substance $\rightarrow PbO_2$

Oxidising agent $\rightarrow PbO$

14. Consider the reactions:

$$2S_2O_3^{2-}(aq)+I_2(s)\rightarrow S_4O_6^{2-}(aq)+2I^-(aq)$$

$$S_2O_3^{2-}(aq)+2Br_2(l)+5H_2O(l)\rightarrow 2SO_4^{2-}(aq)+4Br^-(aq)+10H^+(aq)$$

Why does the same reductant, thiosulphate react differently with iodine and bromine?

Ans: $$2S_2O_3^{2-}(aq)+I_2(s)\rightarrow S_4O_6^{2-}(aq)+2I^-(aq)$$

$$S_2O_3^{2-}(aq)+2\overset{+2\ -2}{Br_2}(l)+5H_2O(l)\rightarrow$$

$$2SO_4^{+6\ 2-}(aq)+4Br^-(aq)+10H^+(aq)$$

When compared to I_2, bromine is a more powerful oxidizer. In SO_4^{2-}, it oxidises the S of $S_2O_3^{2}$ to a higher oxidation state $+6$.

In $S_4O_6^{2-}, I_2$ oxidises S from $S_2O_3^{2}$ to a lower oxidatión state of 2.5 . As a result, the same reductant, thiosulphate, reacts with bromine and iodine in distinct ways.

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

Ans: F_2 can also oxidize Cl to Cl_2, Br^- to Br_2 and I^- to I_2

$$F_{2(aq)}+2Cl_{(s)}\rightarrow 2F_{(aq)}+Cl_{(g)}$$

$$P_{2(aq)}+2I_{(aq)}\rightarrow 2F_{(aq)}+I_{2(s)}$$

Cl_2, Br_2, and I_2, on the other hand, are unable to convert F^- to F_2. Halogens have an oxidizing power of $I_2 < Br_2 < Cl_2 < F_2$. Fluorine, as a result, is the best halogen oxidant.

H_2SO_4 can be converted to SO_2 using HI and HBr, but not with HCl or HF. HI and HBr are thus more effective reductants than HCl and HF

$$2HI + H_2SO_4 \rightarrow I_2 + SO_2 + 2H_2O$$

$$2HBr + H_2SO_4 \rightarrow Br_2 + SO_2 + 2H_2O$$

I can reduce Cu^{2+} to Cu^+ once more, whereas Br^- cannot.

$$4I_{(aq)} + 2Cu^{2+}(aq) \rightarrow Cu_2I_{2(f)} + I_{2(aq)}$$

As a result, among hydrohalic compounds, hydroiodic acid is the best reductant.

Hydrohalic acids' reducing power thus grows in the order of $HF < HCl < HBr < HI$

16. Why does the following reaction occur?

$$XeO_6^{4-}(aq) + 2F^-(aq) \rightarrow XeO_3(g) + F_2(g) + 3H_2O(l)$$

What conclusion about the compound $NaXeO_6$ (of which XeO_6^{4} is a part) can be drawn from the reaction?

Ans: Because XeO_6^{4-} oxidizes F^- and F^- decreases XeO_6^{4-}, the stated reaction occurs.

$$XeO_6^{4-}(aq) + 2F^-(aq) \rightarrow XeO_3(g) + F_2(g) + 3H_2O(l)$$

Xe 's oxidation number (O.N.) falls from $+8$ in XeO_6^{4} to $+6$ in XeO_3, while F 's $O.N$. rises from -1 in F^- to O in F_2

As a result, we can deduce that $NaXeO_6^{4}$ is a more powerful oxidizer than F

17. 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) $H_3PO_2(aq) + 2CuSO_4(aq) + 2H_2O(l) \rightarrow H_3PO_4(aq) + 2Cu(s) + H_2SO_4(aq)$

(c) $C_6H_5CHO(l) + 2\left[Ag\left(NH_3\right)_2\right]^{+}(aq) + 3OH^{-}(aq) \rightarrow C_6H_5COO^{-}(aq) + 2Ag(s) + 4NH_3(aq)$

(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?

Ans: In reactions (a) and (b), Ag^{+} and Cu^{2+}, respectively, act as oxidizing agents.

Ag+ oxidizes C_6H_5CHO to $C_6H_5COO^{-}$ in reaction (c), but Cu^{2+} cannot oxidize C_6H_5CHO in reaction (d)

As a result, $Ag+$ is a more powerful oxidizing agent than Cu^{2+}

18. Balance the following redox reactions by ion-electron method:

(a) $MnO^{-}(aq) + I^{-}(aq) \rightarrow MnO_2(s) + I_2(s)$

Ans: Step 1: The following are the two half reactions involved in the given reaction:

Half-reaction of oxidation $I_{(a\phi)} \rightarrow I_{2(s)}$

Half-reaction of reduction MnO_4 (aq) $\rightarrow MnO_{2(aq)}$

Step 2: We have the following equation for balancing I in the oxidation half reaction:

$2F_{(aq)} \rightarrow I_{2(s)}$

To balance the charge, we add $2e^{-}$ to the reaction's RHS.

$2I^{-}_{(aq)} \rightarrow I_{2(s)} + 2e^{-}$

Step 3: Mn 's oxidation state has decreased from+7 to+4 throughout the reduction half reaction.

$$MnO_4^- + 3e^- \rightarrow MnO_{2(aq)}$$

Step 4: Six O atoms are on the RHS and four O atoms are on the LHS in this equation. As a result, the LHS is given two water molecules.

$$MnO^-_{4(aq)} + 2H_2O + 3e^- \rightarrow MnO_{2(aq)} + 4OH$$

Step 5: By multiplying the oxidation half reaction by 3 and the reduction half reaction by 2 , we may equalize the quantity of electrons.

$$6I^-_{(aq)} \rightarrow 3I_{2(s)} + 6e^-$$

$$2MnO_4^-(aq) + 4H_2O + 6e^- \rightarrow 2MnO_{2(aq)} + 8OH^-_{(aq)}$$

Step 6: When the two half reactions are added together, we get the net balanced redox reaction:

$$6I^-_{(aq)} + 2MnO_{4(aq)} + 4H_2O \rightarrow 2MnO_{2(aq)} + 8OH^-_{(aq)}$$

(b) $MnO^-(aq) + SO_2(aq) \rightarrow Mn^{2+}(aq) + HSO_4^-(aq)$

Ans: If we repeat the processes from part (a), we get the following oxidation half reaction:

$$SO_{2(g)} + 2H_2O_{(l)} \rightarrow HSO_4^-(aq) + 3H^+_{(aq)}$$

And the half-reduction reaction is as follows:

We get the net balanced redox reaction by multiplying the oxidation half reaction by 5 and the reduction half reaction by 2 , then adding them.

(c) $H_2O_2(aq) + Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + H_2O(l)$

Ans: Using the same techniques as in part (a), we get the following oxidation half reaction:

$$Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e$$

And the half-reduction reaction is as follows:

$$H_2O_{2(aq)} + 2H^+_{(aq)} + 2e^- \rightarrow 2H_2O_{(i)}$$

We get the net balanced redox reaction by multiplying the oxidation half reaction by 2 and then adding it to the reduction half reaction:

$$H_2O_{2(aq)} + 2Fe^{2+}_{(aq)} + 2H^+_{(ac)} \rightarrow 2Fe^{3+}_{(2q)} + 2H_2O_{(i)}$$

(d) $Cr_2O_7^{2-} + SO_2(g) \rightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$

Ans: Using the same techniques as in part (a), we get the following oxidation half reaction:

$$SO_{2(g)} + 2H_2O_{(l)} \rightarrow SO^{2-}_{4(2q)} + 4H^+_{(aq)} + 2e^-$$

And the half-reduction reaction is as follows:

$$Cr_2O^{2-}_{7(aq)} + 14H^+_{(aq)} + 6e^- \rightarrow 2Cr^{3+}_{(aq)} + 7H_2O_{(i)}$$

We get the net balanced redox reaction by multiplying the oxidation half reaction by 3 and then adding it to the reduction half reaction:

$$Cr_2O^{2-}_{7(aq)} + 3SO_{2(g)} + 2H^+_{(aq)} \rightarrow 2Cr^{3+}_{(a_9)} + 3SO^{2-}_{4(aq)} + H_2O_{(l)}$$

19. Balance the following equations in basic medium by ion-electron method and oxidation number methods and identify the oxidizing agent and the reducing agent.

$$P_{4(s)} + OH^-_{(aq)} \rightarrow PH_{3g} + HPO^-_{2(aq)}$$

Ans: (a) The oxidation number of P drops from 0 to -3 in P_4 and increases from 0 to $+2$ in HPO_2. As a result, P_4 serves as both an oxidizing and reducing agent in this process. Ion-electron method:

The half-equation for oxidation is:

$$P_{4(s)} \rightarrow HPO^-_{2(aq)}$$

The P atom is balanced in the following way:

$P^0_{4(s)} \rightarrow 4HPO^-_{2(aq)}$

The O.N. is balanced by adding eight electrons in the following way:

$P_{4(s)} \rightarrow 4HPO_{2(a)} + 8e^-$

The charge is balanced by the addition of $12OH^-$ as follows:

$P_{4(s)} + 12OH^-_{(aq)} \rightarrow 4HPO^-_{2(aq)} + 8e$

By adding $4H_2O$, the H and O atoms are balanced.

$P_{4(s)} + 12OH^-_{(aq)} \rightarrow 4HPO^-_{2(aq)} + 4H_2O_{(1)} + 8e^-$ ..(i)

The half-reduction equation is as follows:

$P_{4(s)} + PH_{3(g)}$

The P atom is in a state of equilibrium.

By adding 12 electrons to the equation., it is balanced:

$P_{4(s)} + 12e^- \rightarrow 4PH_{3(g)}$

The charge is balanced by the addition of $12OH^-$ as follows:

$P_{4(s)} + 12e^- \rightarrow 4PH_{3(g)} + 12OH^-_{(aq)}$

$12H_2O$ is used to balance the 0 and H atoms as follows:

$P_{4(s)} + 12H_3O_{(1)} + 12e^- \rightarrow 4PH_{3(g)} + 12OH^-_{(2q)}$ (ii)

The balanced chemical equation can be found by multiplying equations I and (ii) by 3 and then adding them.

$5P_{4(s)} + 12H_2O_{(i)} + 12HO^-_{(aq)} \rightarrow 8PH_{3(g)} + 12HPO^-_{(aq)}$

(b) $N_2H_4(l) + ClO_3^-(aq) \rightarrow NO(g) + Cl^-(g)$

Ans: N' s oxidation number rises from -2 in N_2H_4 to $+2$ in NO, while Cl 's oxidation number falls from $+5$ in ClO_3 to -1 in Cl^-.As a result, N_2H_4 is the reducing agent and ClO_3 is the oxidizing agent in this reaction.

Ion-electron method:

The half-equation for oxidation is:

-2

$N_2H_{4(1)} \rightarrow NO_{(g)}$

The N atoms are balanced in the following way:

$N_2H_{4(1)} \rightarrow 2NO_{(g)}$

By adding 8 electrons to the oxidation number, the oxidation number is balanced:

$N_2H_{4(l)} \rightarrow 2NO_{fg)} + 8e^-$

$8OH^-$ ions are added to balance the charge as follows:

$N_2H_{4(I)} + 8OH^-_{(aq)} \rightarrow 2NO_{(g)} + 8e^-$

$6H_2O$ is added to balance the O atoms as follows:

$N_2H_{4(1)} + 8OH^-_{(aq)} \rightarrow 2NO_{(g)} + 6H_2O \quad 8e^-$.

The half-reduction equation is as follows:

$\overset{+5}{Cl}O^-{}_{3(aq)} \rightarrow \overset{-1}{Cl}{}^-_{(aq)}$

By adding 6 electrons to the oxidation number, the oxidation number is balanced:

$ClO_{3(aq)} + 6e^- \rightarrow Cl^-_{(aq)}$

$6OH^-$ ions are added to balance the charge as follows:

$ClO^-_{3(aq)} + 6e^- \rightarrow Cl^-_{(aq)} + 6OH^-_{(aq)}$

By adding $3H_2O$ as follows, the O atoms are balanced.

Equation I is multiplied by 3 and equation (ii) is multiplied by 4, resulting in the balanced equation:

$$3N_2H_{4(l)} \rightarrow 6NO_{(g)} + 4Cl^-_{(aq)} + 6H_2O_{(l)}$$

Oxidation number method:

Total reduction in N oxidation number

$$N = 2 \times 4 = 8$$

Total reduction in Cl oxidation number

$$Cl = 1 \times 6 = 6$$

To balance the rise and decrease in O.N., multiply N_2H_4 by three and ClO_3 by four.

$$3N_2H_{4(l)} \rightarrow 4ClO^-_{3(aq)} \rightarrow NO_{(g)} + Cl^-_{(aq)}$$

The atoms of N and Cl are balanced as follows:

$$3N_2H_{4(l)} \rightarrow 4ClO^-_{3(aq)} \rightarrow 6NO_{(g)} + 4Cl^{\circ}_{(aq)}$$

$6H_2O$ is added to balance the O atoms as follows:

$$3N_2H_{4(1)} \rightarrow 4ClO_{3(aq)} \rightarrow 6NO_{(g)} + 4Cl^-_{(aq)} + 6H_2O$$

This is the equation that must be balanced.

(c) $Cl_2O_{7(g)} + H_2O_{2(\alpha q)} \rightarrow ClO^-_{2(\alpha q)} + O_{2(g)} + H^+_{(c)}$

Ans: The oxidation number of Cl decreases from $+7$ in Cl_2O_7 to $+3$ in ClO_2 and the oxidation number of O increases from -1 in H_2O_2 to zero in O_2. Hence, in this reaction, Cl_2O_7 is the oxidizing agent and H_2O_2 is the reducing agent. Ion-electron method:

The half-equation for oxidation is:

$$H_2O_{2(aq)} \rightarrow O^0_{2(g)}$$

By adding two electrons to the oxidation number, the oxidation number is balanced as follows:

$H_2O_{2(aq)} \rightarrow O_{2(g)} + 2e^-$

$2OH$ - ions are added to balance the charge as follows:

$H_2O_{2(aq)} + 2OH^-_{2(g)} \rightarrow O_{2(g)} + 2e^-$

By adding $2H_2O_2$ as follows, the oxygen atoms are balanced.

$H_2O_{2(aq)} + 2OH^-_{2(g)} \rightarrow O_{2(g)} + 2H_2O_{2(t)} + 2e^-$.(i).

The half-reduction equation is as follows:

The Cl atoms are balanced in the following way:

$Cl_2O_{7(g)} \rightarrow ClO^-_{2(aq)}$

By adding 8 electrons to the oxidation number, the oxidation number is balanced:

$Cl_2O_{7(g)} + 8e^- \rightarrow 2ClO^-_{2(aq)}$

6OH' is added to balance the charge as follows:

$Cl_2O_{7(g)} + 8e^- \rightarrow 2ClO^-_{2(aq)} + 6OH^-_{(aq)}$

By adding $3H_2O$ as follows, the oxygen atoms are balanced.

$Cl_2O_{7(g)} + 3H_2O + 8e^- \rightarrow 2ClO^-_{2(aq)} + 6OH^-_{(aq)}$. (ii)

By multiplying equation (i) by 4 and adding equation (ii) toit, you can get the balanced equation.

$Cl_2O_{7(g)} + 4H_2O_{2(aq)} + 2OH^-_{(aq)} \rightarrow 2ClO^-_{2(aq)} + 4O_{2(g)} + 5H_2O_{(c)}$

Method for calculating the oxidation number:

The total number of oxidations has decreased $Cl_2O_7 = 4 \times 2 = 8$

The total number of oxidations has decreased $H_2O_2 = 2 \times 1 = 2$

To balance the rise and decrease in the oxidation number, multiply H_2O_2 and O_2 by 4

$Cl_2O_{7Vg} + 4H_2O_{2(aq)} \rightarrow ClO^-_{2(aq)} + 4O_{2(g)}$

The Cl atoms are balanced in the following way:

$Cl_2O_{7(g)} + 4H_2O_{2(aq)} \rightarrow 2ClO^-_{2(aq)} + 4O_{2(g)}$

The O atoms are balanced by adding $3H_2O$ in the following way:

$Cl_2O_{7(g)} + 4H_2O_{2(aq)} \rightarrow 2ClO^-_{2(aq)} + 4O_{2ge} + 3H_2O_{41}$

$2OH^-$ and $2H_2O$ are used to balance the H atoms as follows:

$Cl_2O_{7(g)} + 4H_2O_{2(aq)} + 2OH^- \rightarrow 2ClO^-_{2(aq)} + 4O_{2(g)} + 5H_2O_{(l)}$

This is the equation that must be balanced.

20. What sorts of information can you draw from the following reaction?

$(CN)_{2(g)} + 2OH^-_{(aa)} \rightarrow CN^-_{(2aq)} + CNO^-_{(aq)} + H_2O_{(i)}$

Ans: The carbon oxidation numbers in $(CN)_2, CN^-, CNO^-$ are $+3$, $+2$ and +4respectively.

These can be found as follows:

Let x be C 's oxidation number.

$(CN)_2$

$2(x-3) = 0$

$\therefore x = 3$

CN°

$x-3 = -1$

$\therefore x = 2$

CNO^-

$x - 3 - 2 = -1$

$\therefore x = 4$

The carbon oxidation number in various species is:

$$[CN]_{2(g)} + 2OH^{-}_{(aq)} \rightarrow CN^{*}_{(aq)} + CNO^{-}_{(aq)} + H_2O_{(b)}$$

In the preceding equation, the same chemical is being reduced and oxidized at the same time. Disproportionation reactions are those in which the same chemical is reduced and oxidized at the same time. As a result, the alkaline breakdown of cyanogen can be considered a disproportionation process.

21. The Mn^{3+} ion is unstable in solution and undergoes disproportionation to give Mn^{2+}, MnO_2, and $H+$ ion. Write a balanced ionic equation for the reaction.

Ans: The following is a representation of the provided reaction:

$$Mn^{3+}{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)} + MnO_{2(s)} + H^{+}_{(aq)}$$

The half-equation for oxidation is:

$$\overset{+3}{Mn^{3+}}{}_{(aq)} \rightarrow \overset{+4}{MnO_{2(s)}}$$

By adding one electron to the oxidation number, the oxidation number is balanced as follows:

$$Mn^{3+}{}_{(aq)} \rightarrow MnO_{2(s)} + e^{-}$$

The charge is balanced by introducing $4H^{+}$ ions in the following way:

$$Mn^{3+}{}_{(aq)} \rightarrow MnO_{2(s)} + 4H^{+} + e^{-}$$

$2H_2O$ molecules are added to balance the O atoms and H^{+} ions as follows:

$$Mn^{3+}_{(aq)} + 2H_2O \rightarrow MnO_{2(s)} + 4H^{+} + e^{-} \quad ..(i)$$

The reduction half equation is as follows:

$$Mn^{3+}{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)}$$

By adding one electron to the oxidation number, the oxidation number is balanced:

$$Mn^{3+}_{(aq)} + e^{-} \rightarrow Mn^{2+}_{(aq)}$$

Combining equations I and (ii) yields the balanced chemical equation:

$$Mn^{3+}_{(aq)} + 2H_2O \rightarrow MnO_{2(s)} + 2Mn^{2+}(aq) + 4H^{+}$$

22. Consider the elements:

Cs, Ne, I and F

(a) Identify the element that exhibits only negative oxidation state.

Ans: F has just a -1 negative oxidation state.

(b) Identify the element that exhibits only positive oxidation state.

Ans: Cs **has a positive oxidation state of** $+1$

(c) Identify the element that exhibits both positive and negative oxidation states.

Ans: Both positive and negative oxidation states are present in my body. It has the following oxidation states: -1, $+1$, $+3$, $+5$, and $+7$

(d) Identify the element which exhibits neither the negative nor does the positive oxidation state.

Ans: Ne has a zero-oxidation state. It doesn't have any oxidation states, either negative or positive.

23. Chlorine is used to purify drinking water. Excess of chlorine is harmful. The excess of chlorine is removed by treating with Sulphur dioxide. Present a balanced equation for this redox change taking place in water.

Ans: The following is a representation of the provided redox reaction:

$$Cl_{2(s)} + SO_{2(aq)} + H_2O \rightarrow Cl^{-}_{(aq)} + SO^{2-}_{4(aq)}$$

The half-reaction of oxidation is:

$+4$

$SO_{2(aq)} \rightarrow SO^{2-}$

$4(aq)$

By adding two electrons to the oxidation number, the oxidation number is balanced:

$SO_{2(aq)} \rightarrow SO^{2-}{}_{4(aq)} + 2e^{-}$

The charge is balanced by introducing $4H^{+}$ ions in the following way:

$SO_{2(aq)} \rightarrow SO^{2-}{}_{4(aq)} + 4H^{+} + 2e$

$2H_2O$ molecules are added to balance the O atoms and $H+$ ions as follows: The charge is balanced by introducing $4H^{+}$ ions in the following way:

$SO_{2(aq)} + 2H_2O_{(l)} \rightarrow SO^{2-}{}_{4(aq)} + 4H^{+} + 2e$

The half-reduction reaction is as follows:

$Cl_{2(s)} \rightarrow Cl^{-}_{(aq)}$

The chlorine atoms are balanced in the following way:

$\overset{0}{Cl_{2(s)}} \rightarrow \overset{-1}{Cl^{-}}(aq)$

By adding electrons, the oxidation number is restored.

$Cl_{2(s)} + 2e^{-} \rightarrow 2Cl^{-}_{(aq)} \ldots (ii)$

Combining equations I and (ii) yields the balanced chemical equation:

$Cl_{2(s)} + 2SO_{2(aq)} + 2H_2O_{(i)} \rightarrow 2Cl^{-}_{(aq)} + SO^{2-}{}_{4(aq)} + 4H^{+}_{(aq)}$

24. Refer to the periodic table given in your book and now Ans: the following questions:

(a) Select the possible non-metals that can show disproportionation reaction.

Ans: One of the reacting compounds must always contain an element that can exist in at least three oxidation states in disproportionation reactions.

(a) Because these elements can exist in three or more oxidation states, disproportionation reactions can occur.

(b) Select three metals that can show a disproportionation reaction.

Ans: Because these elements can exist in three or more oxidation states, disproportionation reactions can occur.

25. In Ostwald's process for the manufacture of nitric acid, the first step involves the oxidation of ammonia gas by oxygen gas to give nitric oxide gas and steam. What is the maximum weight of nitric oxide that can be obtained starting only with 10.00 g. of ammonia and 20.00 g of oxygen ?

Ans: For the above reaction, the balanced chemical equation is:

$$4NH_{3(g)} + 5O_{2(g)} \rightarrow 4NO_{(g)} + 6H_2O_{(g)}$$

$$4\times 17\,g \quad 5\times 32\,g \quad 4\times 30\,g \quad 6\times 18\,g$$

$$= 68\,g \quad = 160\,g \quad = 120\,g \quad = 108\,g$$

Therefore, $68\,g$ of NH_3 reacts with $160\,g$ of O_2

Thus, $10\,g$ of NH_3 reacts with

$$\frac{160\times 10}{68}\,g \text{ of } O_2$$

$$23.53\,g \text{ of } O_2$$

However, there is only $20\,g$ of oxygen accessible.

As a result, O_2 is the reaction limiting reagent (we used the amount of O2 to compute the weight of nitric oxide produced).

Hence, $160\,g$ of O_2 gives $120\,g$ of NO

$20\,g$ of O_2 gives $\frac{120 \times 20}{160} g$ of N

Thus, $15\,g$ of NO

As a result, you can get up to $15\,g$ of nitric oxide.

26. Using the standard electrode potentials given in the Table 8.1, predict if the reaction between the following is feasible:

(a) $Fe^{3}(aq)$ **and** $I^{-}(aq)$

Ans: The reaction between $Fe^{3+}(aq)$ and $I^{-}(aq)$ can be expressed as,

$$2Fe^{3+}_{(aq)} + 2I^{-} \rightarrow 2Fe^{2+}_{(aq)} + I_{2(s)}$$

Half-equation for oxidation: $2I^{-} \rightarrow I_{2(s)} + 2e^{-}; E^{0} = -0.54\,V$ Half-equation of reduction

$$\frac{\left[Fe^{3+}(aq) + e^{-} \rightarrow Fe^{2+}_{(aq)}\right] \times 2; E^{0} = +0.77\,V}{2Fe^{3+}_{(aq)} + 2I^{-} \rightarrow 2Fe^{2+}_{(aq)} + I_{2(s)}; E^{0} = +0.23\,V}$$

The overall reaction has an E^{0} of favourable. As a result, the reaction of Fe^{3+} and $I^{-}_{(aq)}$ is possible.

(b) $Ag^{+}(aq)$ **and** $Cu(s)$

Ans: The reaction between $Ag^{+}(aq)$ and $Cu(s)$ can be described as follows:

$$2Ag^{+}_{(aq)} + Cu_{(s)} \rightarrow 2Ag_{(s)} + Cu^{2+}$$

Half-equation for oxidation: $Cu_{(s)} \rightarrow Cu^{2+}_{(a)} + 2e^{-}; E^{0} = -0.34\,V$

Half-equation of reduction

$$\frac{\left[Ag^{+}(aq) \rightarrow e^{-} + Ag_{(s)}\right] \times 2; E^{0} = +0.80\,V}{2Ag^{+}_{(aq)} + Cu_{(s)} \rightarrow 2Ag_{(s)} + Cu^{2+}; E^{0} = +0.46\,V}$$

The overall reaction has an E^0 of favourable. As a result, the reaction of $Ag^+(aq)$ and $Cu(s)$ is possible.

(c) $Fe^{3+}(aq)$ **and** $Cu(s)$

Ans: The reaction between $Fe^{3+}(aq)$ and $Cu(s)$ can be described as follows:

$$2Fe^{3+}_{(aq)} + Ca_V \rightarrow Fe^{2+}_{(s)} + Cu^{2+}_{(2q)}$$

Half-equation for oxidation: $Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-; E^0 = 0.34\,V$

Half-equation of reduction

$$\frac{\left[Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(s)}\right] \times 2; E^0 = 0.77\,V}{2Fe^{3+}_{(aq)} + Cu_{(s)} \rightarrow Fe^{2+}_{(s)} + Cu^{2+}_{(aq)}; E^0 = 0.43\,V}$$

The overall reaction has an E^0 is favourable. As a result, the reaction of $Fe^{3+}(aq)$ and $Cu(s)$ is possible.

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

Ans: The reaction between $Ag(s)$ and $Fe^{3+}(aq)$ can be described as follows:

$$Ag_{(s)} + 2Fe_{(aq)} \rightarrow Ag^+_{(aq)} + Fe^{2+}_{(aq)}$$

Half-equation for oxidation: $Ag_{(s)} + \rightarrow Ag^+_{(aq)} + e^-; E^0 = 0.80\,V$

Half-equation of reduction

$$\frac{\left[Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(s)}\right] \times 2; E^0 = 0.77\,V}{Ag_{(s)} + 2Fe^{3+}_{(aq)} \rightarrow Ag^+_{(aq)} + Fe^{2+}_{(aq)}; E^0 = 0.03\,V}$$

The overall reaction has an E^0 is not favourable. As a result, the reaction of $Ag(s)$ and $Fe^{3+}(aq)$ is not possible.

(e) $Br_{2(aq)}$ **and** $Fe^{2+}(aq)$

Ans: The reaction between $Br_{2(aq)}$ and $Fe^{2+}(aq)$ can be described as follows

Half-equation for oxidation: $Br_{2(s)} + 2Fe^{2+}_{(aq)} \rightarrow 2Br^-_{(aq)} + 2Fe^{3+}_{(aq)}$

Half-equation of reduction: $\left[Fe^{2+}_{(aq)} \rightarrow Fe^{3+}(aq) + e^{-}\right] \times 2; E^0 = -0.77\,V$

Half-equation of reduction

$$\frac{Br_{2(aq)} + 2e^- \rightarrow 2Br^-_{(aq)}; E^0 = +1.09\,V}{Br_{2(aq)} + 2Fe^{2+}_{(aq)} \rightarrow 2Br^-_{(aq)} + 2Fe^{3+}_{(aq)}; E^0 = +0.32\,V}$$

27. Predict the products of electrolysis in each of the following:

(i) An aqueous solution of $AgNO_3$ with silver electrodes

Ans: The reaction between Ag and Cu can be described as follows:

At the cathode, electrolysis can decrease either Ag^+ ions or H_2O molecules. However, Ag^+ ions have a larger reduction potential than H_2O

$$Ag_{(aqi} + e^- \rightarrow Ag_{(s)}; E^0 = +0.80\,V$$

$$2H_2O_{(l)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(2q)}; E^0 = -0.83\,V$$

As a result, at the cathode, Ag^+ ions are decreased. At the anode, Ag metal or H_2O molecules can also be oxidised. However, Ag molecules have a larger oxidation potential than H_2O molecules.

$$Ag_{(s)} \rightarrow Ag^+_{(aq)} + e^-; E^0 = -0.80\,V$$

$$2H_2O_{(i)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}; E^0 = -0.83\,V$$

As a result, Ag metal oxidises at the anode.

(ii) An aqueous solution $AgNO_3$ with platinum electrodes

Ans: Pt is difficult to oxidise. As a result, at the anode, water is oxidised, releasing O_2. Ag^+ ions are reduced and deposited at the cathode.

In aqueous solutions, H_2SO_4 ionises to give

iii. A dilute solution of H_{2S} and SO^2_4 ions.

$H_2SO_{4(aq)} \rightarrow 2H^+_{(aq} + SO^{2-}_{4(aq)}$

At the cathode, electrolysis can decrease either H^+ ions or H_2O molecules. H^+ ions, on the other hand, have a greater reduction potential than H_2O molecules.

$$2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}; E^0 = 0.0\,V$$

$$2H_2O_{(aq)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}; E^0 = -0.83\,V$$

As a result, H^+ ions are reduced at the cathode, releasing H_2 gas.

The anode, on the other hand, can oxidise either two SO_4 ions or two H_2O molecules.

However, when $SO^{2-}{}_4$ is oxidised, more bonds are broken than when H_2O molecules are oxidised.

As a result, the oxidation potential of $SO^{2-}{}_4$ ions is lower than that of H_2O. As a result, H_2O is oxidised at the anode, releasing O_2 molecules.

(iv)An aqueous solution of $CuCl_2$ with platinum electrodes.

Ans: $CuCl_2$ ionises in aqueous solutions to produce Cu^{2+} and Cl^- ions as Cu^{2+} and Cl^- ions.

$$CuCl_{2(aq)} \rightarrow Cu^{2+}{}_{(aq)} + 2Cl^-$$

Cu^{2+} ions or H_2O molecules can be reduced at the cathode during electrolysis. Cu^{2+} on the other hand, has a greater reduction potential than H_2O molecules.

Similarly, either Cl^- or H_2O is oxidised at the anode. The oxidation potential of H_2O is greater than the oxidation potential of Cl^-.

$$2Cl^*_{(aq)} \rightarrow Cl_{2(g)} + 2e^-; E^0 = -1.36\,V$$

$$2H_2O_{(1)} \rightarrow O_{2(g)} + 4H^+_{(2q)} + 4e^-; E^0 = -1.23\,V$$

However, due to over-voltage, oxidation of H_2O molecules occurs at a lower electrode potential than that of Cl^- ions (extra voltage required to liberate gas). As a result, at the anode, Cl^- ions are oxidised, releasing Cl_2 gas.

28. Arrange the following metals in the order in which they displace each other from the solution of their salts.

Al, Cu, Fe, Mg and Zh

Ans: A metal with a higher reducing power displaces a metal with a lower reducing power from its salt solution.

Al, Cu, Fe, Mg and Zn are the metals in order of increasing reducing power. As a result, we can conclude that Mg can evict Al from its salt solution, while Al cannot evict Mg. As a result, the following is the sequence in which the supplied metals displace each other from the solution of respective salts:

$Mg > Al > Zn > Fe > Cu$

29. Given the standard electrode potentials,

$K^+ / K = -2.93\,V, Ag^+ / Ag = 0.80\,V$

$Hg^+ / Hg - 0.79$

$Mg^{2+} / Mg = -2.37\,V\ \ Cr^{3+} / Cr = -0.74\,V$

Arrange these metals in their increasing order of reducing power.

Ans: The stronger the reducing agent is, the lower the electrode potential. As a result, the reducing power of the above metals is in ascending order: $Ag > Hg > Cr > Mg > K$

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

takes place, further show

(i) which of the electrode is negatively charged,

(ii) the carriers of the current in the cell, and

(iii) individual reaction at each electrode.

Ans: The galvanic cell that corresponds to the given redox reaction looks like this:

$$Zn\left[Zn^{2+}_{(aq)} \| Ag^{+}_{(aq)}\right]Ag$$

(i) Because Zn oxidises to Zn^{2+} at this electrode, the remaining electrons concentrate on it, the Zn electrode is negatively charged.

(ii) Ions are the current carriers in cells.

(iii) The reaction at the Zn electrode can be represented as follows:

$$Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2e^{-}$$

The reaction at the Ag electrode can be represented as follows:

$$Ag^{+}{}_{(aq)} + e^{\circ} \rightarrow Ag_{(s)}$$

(iv) $CuCl_2$ ionises in aqueous solutions to produce Cu^{2+} and Cl^{-} ions as:

$$CuCl_{2(aq)} \rightarrow Cu^{2+}{}_{(aq)} + 2Cl^{-}{}_{(aq)}$$

Cu^{2+} ions or H_2O molecules can be reduced at the cathode during electrolysis.

$$Cu^{2+}{}_{(aq)} + 2e^{-} \rightarrow Cu_{(aq)}; E^0 = 0.34\,V$$

$$H_2O_{(i)} + 2e^{-} \rightarrow H_{2(g)} + 2OH^{-}; E^0 = -0.83\,V$$

Cu^{2+} ions are so reduced and deposited at the cathode.

The oxidation potential of H_2O is greater than the oxidation potential of Cl^{-}.

$$2Cl^{-}_{(aq)} + Cl_{2(H)} \rightarrow 2e^{-}; E^0 = 1.36\,V$$

$$2H_2O_{(i)} + O_{2(g)} \rightarrow 4H^{+}{}_{2(g)} + 4e^{-}; E^0 = 1.23\,V$$

However, due to over-voltage, oxidation of H_2O molecules occurs at a lower electrode potential than that of Cl^{-} ions (extra voltage required to liberate gas).

30. **Depict the galvanic cell in which the reaction**

$Zn(s) + 2Ag^{+}(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$ takes place, Further show:

(i) Which of the electrode is negatively charged?

(ii) The carriers of the current in the cell.

(iii) individual reaction at each electrode.

Ans: The galvanic cell in which the given reaction takes place is depicted as:

$Zn(s)|Zn^{2+}(aq)||Ag^{+}(aq)|Ag(s)$

(i) The negatively charged electrode is the zinc electrode. It acts as an anode.

(ii) In the external circuit, the current will flow from silver to zinc.

(iii) Oxidation at anode: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-}$

Reduction at cathode: $Ag^{+}(aq) + e^{-} \longrightarrow Ag(s)$