#423452
Topic: Oxidation Number

Passage
Assign oxidation number to the underlined elements in each of the following species.

NaH$_2$PO$_4$

Solution
Let X be the oxidation state of P in NaH$_2$PO$_4$.

$+1 + 2(1) + X + 4(-2) = +5$

Hence, the oxidation state of P in NaH$_2$PO$_4$ is +5.

#423454
Topic: Oxidation Number

Passage
Assign oxidation number to the underlined elements in each of the following species.

NaHSO$_4$

Solution
Let X be the oxidation state of S in NaHSO$_4$.

$+1 + 1 + X + 4(2) = 0$

$X = +6$

Hence, the oxidation state of S in NaHSO$_4$ is +6.

#423455
Topic: Oxidation Number

Passage
Assign oxidation number to the underlined elements in each of the following species.

H$_4$P$_2$O$_7$

Solution
Let X be the oxidation state of P in H$_4$P$_2$O$_7$.

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

$X = +5$

Hence, the oxidation state of P in H$_4$P$_2$O$_7$ is +5.

#423464
Topic: Oxidation Number

Passage
Assign oxidation number to the underlined elements in each of the following species.

K$_2$MnO$_4$

Solution
Let $x$ be the oxidation number of Mn in $K_2MnO_4$.

\[
2(\, + 1\, ) + X + 4(\, - 2\, ) = 0 \\
2 + X - 8 = 0 \\
X = 6
\]

---

**#423468**  
**Topic:** Oxidation Number  

Assign oxidation number to the underlined elements in each of the following species.

\[Ca^O\]  
\[-2\]

**Solution**  

Let $x$ be the oxidation number of O in $CaO_2$.

\[
2 + 2x = 0 \\
x = -1
\]

---

**#423470**  
**Topic:** Oxidation Number  

**Passage**  

Assign oxidation number to the underlined elements in each of the following species.

\[Na^B \ H_4\]

**Solution**  

Let $x$ be the oxidation number of B in $NaBH_4$.

\[-1 + x + 4(\, - 1\, ) = 0 \\
x = +3
\]

---

**#423474**  
**Topic:** Oxidation Number  

**Passage**  

Complete the following reactions:

\[H_2(g) + MnO_2(s) \overset{\Delta}{\rightarrow} \]

**Solution**  

The completed reaction is

\[\alpha H_2(g) + \frac{m}{n} MnO_2(s) \overset{\Delta}{\rightarrow} m\, Mn(\, s\, ) + \alpha H_2O(l)\]

$\alpha$ moles of $H_2$ react with 1 mole of $MnO_2(s)$ to give $m$ moles of $Mn(\, s\, )$ and $\alpha$ moles of $H_2O(l)$

---

**#423477**  
**Topic:** Oxidation Number  

**Passage**  

Assign oxidation number to the underlined elements in each of the following species.

\[H_2S\]  
\[-2\]

**Solution**
Let $X$ be the oxidation state of S in $\text{H}_2\text{S}_2\text{O}_7$.

\[
2(\text{H}^+) + 2X + 7(2^-) = 0
\]
\[
X = +6
\]

Hence, the oxidation state of S in $\text{H}_2\text{S}_2\text{O}_7$ is +6.

---

**#423481**
**Topic:** Oxidation Number

**Passage**

Assign oxidation number to the underlined elements in each of the following species.

$\text{KAI}^5\text{O}_{4,12}\text{H}_2\text{O}$

**Solution**

Let $x$ be the oxidation state of S in $\text{KAI}^5\text{O}_{4,12}\text{H}_2\text{O}$.

\[
+1 + 3 + 2(x + 4(−2)) = 0
\]
\[
+4 + 2x − 16 = 0
\]
\[
2x = 12
\]
\[
x = +6
\]

---

**#423491**
**Topic:** Oxidation Number

What are the oxidation number of the underlined elements in each of the following and how do you rationalise your results?

$\text{KI}_3$

**Solution**

Let $X$ be the oxidation number of I in $\text{KI}_3$.

\[
+1 + 3X = 0
\]
\[
X = \frac{1}{3}
\]

But the oxidation number cannot be fractional.

$\text{KI}_3$ exists as $\text{K}^+[-I + \frac{1}{3}]^-$. A coordinate bond is formed between $I_2$ molecule and $I^-$ ion. The oxidation number of two I atoms in $I_2$ molecule is 0 and that of $I^-$ ion is -1.

---

**#423493**
**Topic:** Oxidation Number

What are the oxidation number of the underlined elements in each of the following and how do you rationalise your results?

$\text{H}_2\text{S}_{4,6}$

**Solution**
Let $X$ be the oxidation state of $S$ in $H_2S_4O_6$.

$$2(\frac{1}{2}) + 4X + 6(2 - 1) = 0$$

$$X = \frac{5}{2}$$

Hence, the oxidation state of $S$ in $H_2S_4O_6$ is $+5/2$.

But oxidation number cannot be fractional.

Terminal $S$ atoms have $+5$ oxidation number and middle $S$ atoms have $0$ oxidation number.
Passage

What are the oxidation number of the underlined elements in each of the following and how do you rationalise your results?

\[
\begin{align*}
\text{CH}_3\text{COOH} \\
- & -
\end{align*}
\]

Solution

(a) By conventional method:

We will determine average oxidation number of C atom.

Let \( x \) be the oxidation number of C in \( \text{CH}_3\text{COOH} \).

The oxidation numbers of H and O are +1 and -2 respectively.

\[
2x + 4(1) + 2(-2) = 0
\]

\( x = 0 \)

Thus, the oxidation number of both the carbon atoms is zero.

(b) According to the structure:

We will determine the oxidation number of each C atom.

(i) For C atom of -COOH group, let \( x \) be the oxidation number of this C atom.

This C is attached to one O atom by double bond, one -OH group and one methyl group. The oxidation number of O atom attached by double bond is -2. The oxidation number of -OH group is -1. The C atom of methyl group will not affect the oxidation number of -COOH group.

\[
x + 1 + 1(-2) + 1(-1) = 0
\]

\( x = 2 \)

(ii) For methyl carbon atom, let \( x \) be the oxidation number of this C atom.

This C atom is attached to 3 H atoms and one -COOH group.

The oxidation number of H atom is +1 and the -COOH group does not affect the oxidation number of C atom of methyl group.

\[
3(1) + x + 1(-1) = 0
\]

\( x = -2 \)

#423513

**Topic:** Types of redox reactions

Passage

Justify that the following reactions are redox reactions:

\[ \text{CuO}(s) + \text{H}_2(g) \rightarrow \text{Cu}(s) + \text{H}_2\text{O}(g) \]

Solution

The oxidation number of Cu decreases from +2 to 0 and that of H increases from 0 to +1. Hence, \( \text{CuO} \) is reduced to Cu and \( \text{H}_2 \) is oxidized to \( \text{H}_2\text{O} \).

Hence, it is a redox reaction.

#423538

**Topic:** Types of redox reactions

Passage

2\( \text{K}(s) + \text{F}_2(g) \rightarrow 2\text{K}^+ \text{F}^- (aq) \) is a type of ______ reaction.

A disproportionation  
B combustion  
C corrosion  
D redox

Solution

The oxidation number of K increases from 0 to +1 and the oxidation number of F2 decreases from 0 to -1. Hence, K is oxidized and \( \text{F}_2 \) is reduced. Hence, it is redox reaction.
Fluorine reacts with ice and results as follows:

\[ \text{H}_2\text{O(s)} + \text{F}_2(g) \rightarrow \text{HF}(g) + \text{HOF(s)} \]

Justify that this reaction is a redox reaction.

**Solution**

The oxidation number of \( \text{F}_2 \) changes from 0 to -1. Thus, it is reduced. The oxidation number of oxygen changes from -2 to 0. Hence, it is oxidized. Thus, it is a redox reaction.

---

### Types of redox reactions

Passage

Complete the following chemical reactions. Classify the below into (a) hydrolysis, (b) redox and (c) hydration reactions.

\[ \text{MnO}_4^- \text{(aq)} + \text{H}_2\text{O}_2 \text{(aq)} \rightarrow \]

**Solution**

The complete chemical equation is given below:

\[ \text{MnO}_4^- \text{(aq)} + 5\text{H}_2\text{O}_2 \text{(aq)} + 6\text{H}^+ \rightarrow 2\text{Mn}^{2+} \text{(aq)} + 8\text{H}_2\text{O(l)} + 5\text{O}_2(g) \]

This is an example of redox reaction. Mn is reduced and H2O2 is oxidized.

---

### Oxidation Number

Calculate the oxidation number of sulphur, chromium and nitrogen in \( \text{H}_2\text{SO}_3 \), \( \text{Cr}_2\text{O}_7^{2-} \) and \( \text{NO}_2^- \). Suggest structure of these compounds. Count for the fallacy.

**Solution**
(a) $H_2SO_5$ by conventional method.

Let $x$ be the oxidation number of $S$.

$$2(\text{+1}) + x + 5(-2) = 0$$

$$x = +8$$

+8 Oxidation state of $S$ is not possible as $S$ cannot have oxidation number more than 6. The fallacy is overcome if we calculate the oxidation number from its structure.

$$HO - S(O_2) - O - O - H$$

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

$$x = +6$$

(b) Dichromate ion

Let $x$ be the oxidation number of $Cr$ in dichromate ion.

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

$$x = +6$$

Hence the oxidation number of $Cr$ in dichromate ion is +6. This is correct and there is no fallacy.

(c) Nitrate ion, by conventional method

Let $x$ be the oxidation number of $N$ in nitrate ion.

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

From the structure $O^- - N^+ (O) - O^-$

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

$$x = +5$$

Thus there is no fallacy.

---

**Sulphuric acid**  
**Dichromate**  
**Nitrate**
Formula for Nickel (II) sulphate is $\text{NiSO}_4$.

---

**#423562**

**Topic:** Oxidation Number

**Passage**

Write formulas for the following compounds.

**Tin (IV) oxide**

**Solution**

As Tin has $+4$ oxidation state, and oxidation state of $\text{O}$ is $-2$.

So, there must be $2\text{ O}$ atom so that neutral compound can be formed with correct formula.

So, formula must be $\text{SnO}_2$.

---

**#423565**

**Topic:** Oxidation Number

**Passage**

Write formulas for the following compounds.

**Thallium (I) sulphate**

**Solution**

Formula for thallium(I)sulphate is $\text{Tl}_2\text{SO}_4$.

---

**#423566**

**Topic:** Oxidation Number

**Passage**

Write formulas for the following compounds.

**Iron (III) sulphate**

**Solution**

Formula for Iron (III) sulphate is $\text{Fe}_2(\text{SO}_4)_3$.

---

**#423569**

**Topic:** Oxidation Number

**Passage**

Write formulas for the following compounds.

**Chromium (III) oxide**

**Solution**

Formula for chromium (III) oxide is $\text{Cr}_2\text{O}_3$.

---

**#423571**

**Topic:** Oxidation Number

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

**Solution**
The substances along with oxidation states of C are shown below.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Oxidation number of C</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH₂O₂</td>
<td>0</td>
</tr>
<tr>
<td>OF ≡ CF</td>
<td>+1</td>
</tr>
<tr>
<td>HC ≡ CH</td>
<td>-1</td>
</tr>
<tr>
<td>CH₃Cl₂</td>
<td>+2</td>
</tr>
<tr>
<td>CH₃Cl</td>
<td>-2</td>
</tr>
<tr>
<td>Cl₂C – CO₂</td>
<td>+3</td>
</tr>
<tr>
<td>H₃C – CH₃</td>
<td>-3</td>
</tr>
<tr>
<td>CCl₃, CO₂</td>
<td>+4</td>
</tr>
<tr>
<td>CH₄</td>
<td>-4</td>
</tr>
</tbody>
</table>

The substances along with oxidation states of N are shown below.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Oxidation number of N</th>
</tr>
</thead>
<tbody>
<tr>
<td>N₂</td>
<td>0</td>
</tr>
<tr>
<td>N₂O</td>
<td>+1</td>
</tr>
<tr>
<td>N₂H₂</td>
<td>-1</td>
</tr>
<tr>
<td>NO</td>
<td>+2</td>
</tr>
<tr>
<td>N₂H₄</td>
<td>-2</td>
</tr>
<tr>
<td>N₂O₃</td>
<td>+3</td>
</tr>
<tr>
<td>NH₃</td>
<td>-3</td>
</tr>
<tr>
<td>NO₂</td>
<td>+4</td>
</tr>
<tr>
<td>N₂O₃</td>
<td>+5</td>
</tr>
</tbody>
</table>

#423576

**Topic:** Oxidation and reduction - electron transfer concept

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?

**Solution**

The S atom in SO₂ has +4 oxidation number. The minimum and maximum oxidation numbers of S are -2 and +6 respectively. Hence, in SO₂, S can increase and decrease its oxidation number. Hence, SO₂ 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.

#423608

**Topic:** Oxidation and reduction - electron transfer concept

The compound AgF₂ is an unstable compound. However, if formed, the compound acts as a very strong oxidising agent. Why?

**Solution**

Ag in AgF₂ has +2 oxidation state which is an unstable oxidation state of Ag. When AgF₂ is formed, silver accepts an electron to form Ag⁺.

Hence, silver is reduced and AgF₂ acts as a very strong oxidizing agent.

#423615

**Topic:** Oxidation Number

Whenever a reaction between an oxidising 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 oxidising agent is in excess. Justify this statement giving three illustrations.

**Solution**

[Further details on the solution would be included here.]

https://community.toppr.com/content/questions/print/?show_answer=1&show_topic=1&show_solution=1&page=1&qid=423701%2C+423615%2C+4236…
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. Following illustrations justify this.

(i) Oxidizing agent is \( F_2 \) and reducing agent is \( P_4 \). When excess \( P_4 \) reacts with \( F_2 \), \( PF_3 \) is produced in which P has +3 oxidation number.

\[
P_4(\text{excess}) + F_2 \rightarrow PF_3
\]

But if fluorine is in excess, \( PF_5 \) is formed in which P has oxidation number of +5.

\[
P_4 + F_2(\text{excess}) \rightarrow PF_5
\]

(ii) Oxidizing agent is oxygen and reducing agent is K. When excess K reacts with oxygen, \( K_2O \) is formed in which oxygen has oxidation number of -2.

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

But if oxygen is in excess, then \( K_2O_2 \) is formed in which O has oxidation number of -1.

\[
2K + O_2(\text{excess}) \rightarrow K_2O_2
\]

(iii) The oxidizing agent is oxygen and the reducing agent is C. When an excess of C reacts with oxygen, \( CO \) is formed in which C has +2 oxidation number.

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

When excess of oxygen is used, \( CO_2 \) is formed in which C has +4 oxidation number.

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

#423663

**Topic:** Oxidation and reduction - electron transfer concept

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

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

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

**Solution**

Bromine is a stronger oxidizing agent. Hence, it oxidizes \( S_2O_3^{2-} \) to \( SO_4^{2-} \).

Iodine is a weaker oxidizing agent. Hence, it oxidizes \( S_2O_3^{2-} \) to \( S_4O_6^{2-} \).

#423664

**Topic:** Oxidation and reduction - electron transfer concept

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

**Solution**
Fluorine oxidizes chloride ion to chlorine, bromide ion to bromine and iodide ion to iodine respectively.

\[ F_2 + 2Cl^- \rightarrow 2F^- + Cl_2 \]
\[ F_2 + 2Br^- \rightarrow 2F^- + Br_2 \]
\[ F_2 + 2I^- \rightarrow 2F^- + I_2 \]

Chlorine oxidizes bromide ion to bromine and iodide ion to iodine.

\[ Cl_2 + Br^- \rightarrow 2Cl^- + Br_2 \]
\[ Cl_2 + I^- \rightarrow 2Cl^- + I_2 \]

Bromine oxidizes iodide ion to iodine.

\[ Br_2 + I^- \rightarrow 2Br^- + I_2 \]

But bromine and chlorine cannot oxidize fluoride to fluorine. Hence, fluorine is the best oxidizing agent amongst the halogens. The decreasing order of the oxidizing power of halogens is \( F_2 > Cl_2 > Br_2 > I_2 \).

HI and HBr can reduce sulphuric acid to sulphur dioxide but \( HCl \) and \( HF \) cannot. Thus, \( HI \) and \( HBr \) are stronger reducing agents than \( HCl \) and \( HF \).

\[ 2HI + H_2SO_4 \rightarrow I_2 + SO_2 + 2H_2O \]
\[ 2HBr + H_2SO_4 \rightarrow Br_2 + SO_2 + 2H_2O \]

Iodide ion can reduce \( Cu(I) \) to \( Cu(I) \) but bromide cannot.

\[ 4I^- + 2Cu^{2+} \rightarrow 2CuI_2 + I_2 \]

Hence, among the hydrohalic compounds, hydroiodic acid is the best reductant. The reducing power of hydrohalic acids is \( HF < HCl < HBr < HI \).

---

#423692
**Topic:** Oxidation Number

**Passage**

Consider the elements Cs, Ne, I and F.

Identify the element that exhibits only negative oxidation state.

**Solution**

Fluorine is the most electronegative element in the periodic table. It exhibits only negative oxidation state of -1.

---

#423693
**Topic:** Oxidation Number

**Passage**

Consider the elements Cs, Ne, I and F.

Identify the element that exhibits only positive oxidation state.

**Solution**

Cs is the most electropositive element in the periodic table. It is an alkali metal and exhibits only positive oxidation state of +1.

---

#423694
**Topic:** Oxidation Number

**Passage**

Consider the elements Cs, Ne, I and F.

Identify the element that exhibits both positive and negative oxidation states.
Solution

Iodine exhibits both positive and negative oxidation states. It exhibits oxidation states -1, +1, +3, +5 and +7.

---

**#423695**

**Topic:** Oxidation Number

**Passage**

Consider the elements Cs, Ne, I and F.

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

**Solution**

The element which exhibits neither the negative nor does the positive oxidation state is Ne. It is a noble gas with oxidation state of zero.

---

**#423701**

**Topic:** Balance redox reactions

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?

**Solution**

The balanced chemical equation is:

\[ 4\text{NH}_3(g) + 5\text{O}_2(g) \rightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g) \]

The molar masses of ammonia and oxygen are 17 g/mol and 32 g/mol respectively.

5 moles (160 g) of oxygen reacts with 4 moles (68 g) of ammonia.

20 g of oxygen will react with \( \frac{68 \times 20}{160} = 8.5 \) g of ammonia. Hence, oxygen is the limiting reagent.

The molar mass of NO is 30 g/mol.

5 moles (160 g) of oxygen will produce 4 moles (120 g) of NO.

20 g of oxygen will produce \( \frac{120 \times 20}{160} = 15 \) g of NO.

---

**#423705**

**Topic:** Electrode potential

Using the standard electrode potentials given in the table, predict if the reaction between the following is possible.

\( \text{Fe}^{3+} (aq) \) and \( I^- (aq) \)

**Solution**

The oxidation half reaction is \( 2I^- (aq) \rightarrow I_2(s) + 2e^- ; E^0 = -0.54 \text{V} \).

The reduction half reaction is \( [\text{Fe}^{3+} (aq) + e^-] \rightarrow 2\text{Fe}^{2+} (aq) + 0.77 \text{V} = +0.77 \text{V} \).

The net cell reaction is \( 2\text{Fe}^{3+} (aq) + 2I^- (aq) \rightarrow 2\text{Fe}^{2+} (aq) + I_2(s) ; E^0 = +23 \text{V} \).

Since, the cell potential is positive, the reaction is feasible.
Using the standard electrode potentials given in the table, predict if the reaction between the following is possible.

\( Ag^+ (aq) \) and \( Cu(s) \)

**Solution**

Oxidation half reaction is \( Cu(s) \rightarrow Cu^{2+} (aq) + 2e^- \); \( E^0 = -0.34 \text{ V} \).

Reduction half reaction is \( [Ag^+ (aq) + e^- \rightarrow Ag(s)] \times 2; E^0 = +0.80 \text{ V} \).

The net cell reaction is \( 2Ag^+ (aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+} \); \( E^0 = +0.46 \text{ V} \).

Since, the cell potential is positive, the reaction is feasible.

---

Using the standard electrode potentials given in the table, predict if the reaction between the following is possible.

\( Fe^{3+} (aq) \) and \( Cu(s) \)

**Solution**

Oxidation half reaction is \( Cu(s) \rightarrow Cu^{2+} (aq) + 2e^- \); \( E^0 = -0.34 \text{ V} \).

Reduction half reaction is \( [Fe^{3+} (aq) + e^- \rightarrow Fe^{2+} (aq)] \times 2; E^0 = +0.77 \text{ V} \).

The net cell reaction is \( 2Fe^{3+} (aq) + Cu(s) \rightarrow 2Fe^{2+} (aq) + Cu^{2+} \); \( E^0 = +0.43 \text{ V} \).

Since, the cell potential is positive, the reaction is feasible.
Using the standard electrode potentials given in the Table, predict if the reaction between the following is feasible.

\( \text{Ag}(s) \) and \( \text{Fe}^{3+}(aq) \)

**Solution**

Oxidation half reaction is \( \text{Ag}(s) + e^{-} \rightarrow \text{Ag}^{+}(aq); \ E^0 = -0.80 \text{V} \).

Reduction half reaction is \( \text{Fe}^{3+}(aq) + e^{-} \rightarrow \text{Fe}^{2+}(aq); \ E^0 = +0.77 \text{V} \).

The net cell reaction is \( 2\text{Fe}^{3+}(aq) + \text{Ag}(s) \rightarrow 2\text{Fe}^{2+}(aq) + \text{Ag}^{+}; \ E^0 = -0.03 \text{V} \).

Since, the cell potential is negative, the reaction is not feasible.

---

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

\( \text{Br}_2(aq) \) and \( \text{Fe}^{2+}(aq) \)

**Solution**

Oxidation half reaction is \( \text{Fe}^{2+}(aq) + e^{-} \rightarrow \text{Fe}^{3+}(aq); \ E^0 = -0.77 \text{V} \).

Reduction half reaction is \( \text{Br}_2(aq) + 2e^{-} \rightarrow 2\text{Br}^{-}(aq); \ E^0 = +1.09 \text{V} \).

The net cell reaction is \( 2\text{Fe}^{2+}(aq) + \text{Br}_2(aq) \rightarrow 2\text{Fe}^{3+}(aq) + 2\text{Br}^{-}; \ E^0 = -0.32 \text{V} \).

Since, the cell potential is negative, the reaction is not feasible.

---

Predict the products of electrolysis in each of the following.

An aqueous solution of \( \text{AgNO}_3 \) with silver electrodes.

**Solution**

---
Reaction in solution:
\[ \text{AgNO}_3 \rightarrow \text{Ag}^+ + \text{NO}_3^- \]
\[ \text{H}_2\text{O} \rightarrow \text{H}^+ + \text{OH}^- \]

Reaction at cathode:
\[ \text{Ag}^+ + e^- \rightarrow \text{Ag} \]

Reaction at anode:
\[ \text{Ag(s)} + \text{NO}_3^- \rightarrow \text{AgNO}_3(\text{aq}) + e^- \]

In aqueous solution, silver nitrate ionizes to silver ions and nitrate ions. At cathode, either silver ions or water molecules can be reduced. Since, silver ion has higher reduction potential than water, silver ions are reduced at cathode. Similarly, at anode, either silver metal or water molecules can be oxidized. Since oxidation potential of silver is higher than that of water molecules, silver is oxidized.

---

#423717

**Topic:** Types of redox reactions

**Passage**

Predict the products of electrolysis in each of the following.

An aqueous solution \( \text{AgNO}_3 \) with platinum electrodes.

**Solution**

The oxidation of \( \text{Pt} \) is not possible. Water is oxidized at anode which liberates oxygen. Silver ions are reduced at cathode and are deposited.

Reaction in solution:
\[ \text{AgNO}_3 \rightarrow \text{Ag}^+ + \text{NO}_3^- \]
\[ \text{H}_2\text{O} \rightarrow \text{H}^+ + \text{OH}^- \]

Reaction at cathode:
\[ \text{Ag}^+ + e^- \rightarrow \text{Ag} \]

Reaction at anode:

Due to platinum electrode, self ionization of water will take place.
\[ \text{H}_2\text{O} \rightarrow 2\text{H}^+ + \frac{1}{2}\text{O}_2 + 2e^- \]

Hence, silver will deposit at cathode and oxygen gas will generate at anode.

---

#423718

**Topic:** Types of redox reactions

**Passage**

Predict the products of electrolysis in each of the following.

A dilute solution of \( \text{H}_2\text{SO}_4 \) with platinum electrodes.

**Solution**
The dissociation of sulphuric acid gives protons and sulphate ions.

At cathode, either hydrogen ions or water molecules can be reduced. Since protons have higher reduction potential than water, hydrogen ions are reduced to hydrogen gas. At anode, either sulphate ions or water molecule can get oxidized. Since, during oxidation of sulphate ions, more bonds are broken than oxidation of water molecules, sulphate ions have lower oxidation potential than water. Hence, water is oxidized at the anode to liberate oxygen molecules.

\[
\begin{align*}
\text{H}_2\text{SO}_4 & \rightarrow 2\text{H}^+ + \text{SO}_4^{2-} \\
\text{H}_2\text{O} & \rightarrow \text{H}^+ + \text{OH}^- \\
\text{Reaction at cathode;} & 2\text{H}^+ + 2e^- \rightarrow \text{H}_2 \\
\text{Reaction at anode;} & \text{H}_2\text{O} \rightarrow \frac{1}{2}\text{O}_2 + 2e^- \\
\text{Due to platinum electrode, self ionization of water will take place.} & \\
\text{H}_2\text{O} & \rightarrow 2\text{H}^+ + \frac{1}{2}\text{O}_2 + 2e^- \\
\text{Hence, hydrogen gas will generate at cathode and oxygen gas will generate at anode.}
\end{align*}
\]

### #423719
**Topic:** Types of redox reactions

**Passage**

Predict the products of electrolysis in each of the following.

**An aqueous solution of CuCl₂ with platinum electrodes.**

**Solution**

When an aqueous solution of CuCl₂ is electrolyzed with platinum electrodes, chlorine is obtained at anode and Cu is deposited at cathode.

\[
\begin{align*}
2\text{Cl}^- & \rightarrow \text{Cl}_2 + 2e^- \\
\text{Cu}^{2+} + 2e^- & \rightarrow \text{Cu}
\end{align*}
\]

### #423720
**Topic:** Electrode potential

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

Al, Cu, Fe, Mg and Zn.

**Solution**

A metal having stronger reducing power displaces another metal having weaker reducing power from its salt solution.

The increasing order of reducing power is \( \text{Cu} < \text{Fe} < \text{Zn} < \text{Al} < \text{Mg} \).

Thus, Mg can displace Al from its salt solution but Al cannot displace Mg.

The order in which the metals displace each other from their salt solutions is \( \text{Mg} > \text{Al} > \text{Zn} > \text{Fe} > \text{Cu} \).

### #423721
**Topic:** Electrode potential

Given the standard electrode potentials of some metals.

\[
\begin{align*}
\text{K}^+ / \text{K} & = 2.93 \text{V} \\
\text{Ag}^+ / \text{Ag} & = 0.80 \text{V} \\
\text{Hg}^{2+} / \text{Hg} & = 0.79 \text{V} \\
\text{Mg}^{2+} / \text{Mg} & = 2.37 \text{V} \text{and} \\
\text{Cr}^{3+} / \text{Cr} & = 0.74 \text{V}
\end{align*}
\]

Arrange these metals in their increasing order of reducing power.

**Solution**

Lower reduction potential corresponds to higher reducing power. The increasing order of the standard reduction potentials is

\[
\text{K}^+ \text{K} < \text{Mg}^{2+} \text{Mg} < \text{Cr}^{3+} \text{Cr} < \text{Hg}^{2+} \text{Hg} < \text{Ag}^+ \text{Ag}
\]

The increasing order of the reducing power is \( \text{Ag} < \text{Hg} < \text{Cr} < \text{Mg} < \text{K} \).
**#423723**
**Topic:** Electrode potential

**Passage**
Depict the galvanic cell in which the reaction $\text{Zn}(s) + 2\text{Ag}^+(aq) \rightarrow \text{Zn}^{2+}(aq) + 2\text{Ag}(s)$ takes place. Further show:

which of the electrode is negatively charged?

**Solution**
For the given redox reaction, the galvanic cell is:

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

Zinc electrode is negatively charged as Zn is oxidized to $\text{Zn}^{2+}$. The electrons released during oxidation accumulate on this electrode.

---

**#423725**
**Topic:** Electrode potential

**Passage**
Depict the galvanic cell in which the reaction $\text{Zn}(s) + 2\text{Ag}^+(aq) \rightarrow \text{Zn}^{2+}(aq) + 2\text{Ag}(s)$ takes place. Further show:

individual reaction at each electrode.

**Solution**
For the given redox reaction, the galvanic cell is:

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

At $\text{Zn}$ electrode, $\text{Zn}$ is oxidized to $\text{Zn}^{2+}$ ions.

At $\text{Ag}$ electrode, $\text{Ag}^+$ is reduced to $\text{Ag}$

---

**#464730**
**Topic:** Types of redox reactions

Food cans are coated with tin and not with zinc because:

A. zinc is costlier than tin
B. zinc has a higher melting point than tin
C. zinc is more reactive than tin
D. zinc is less reactive than tin

**Solution**
Food cans are coated with tin and not with zinc because zinc is above the tin in reactivity series means more reactive than tin and can react with food elements preserved in it.
#423521
Topic: Oxidation Number

Passage
Justify that the following reactions are redox reactions:

$$\text{Fe}_3\text{O}_4 (s) + 3\text{CO}(g) \rightarrow 2\text{Fe} (s) + 3\text{CO}_2 (g)$$

Solution
The oxidation number of iron decreases from +3 to 0. The oxidation number of C increases from +2 to +4. Hence, $\text{Fe}_3\text{O}_4$ is reduced and CO is oxidized. Hence, it is a redox reaction.

#423532
Topic: Oxidation Number

Passage
Justify that the following reactions are redox reactions:

$$4\text{BCl}_3 (g) + 3\text{LiAlH}_4 (g) \rightarrow 2\text{B}_2\text{H}_6 (g) + 3\text{LiCl}(s) + 3\text{AlCl}_3 (s)$$

Solution
The oxidation number of B decreases from +3 to -3. The oxidation number of hydrogen increases from -1 to +1. Hence, $\text{BCl}_3$ is reduced and $\text{LiAlH}_4$ is oxidized. Hence, it is a redox reaction.

#423537
Topic: Oxidation Number

Passage
Justify that the following reactions are redox reactions:

$$2\text{K} (s) + \text{F}_2 (g) \rightarrow 2\text{KF}^-(s)$$

Solution
The oxidation number of K increases from 0 to +1 and the oxidation number of $\text{F}_2$ decreases from 0 to -1. Hence, K is oxidized and $\text{F}_2$ is reduced. Hence, it is redox reaction.

#423542
Topic: Oxidation and reduction - classical concept

Consider the reaction of water with $\text{F}_2$ and suggest in terms of oxidation and reduction which species are oxidised/reduced.

Solution
The balanced chemical equations are given below:

$$2\text{F}_2 + 2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{F}^-$$
$$3\text{F}_2 + 3\text{H}_2\text{O} \rightarrow \text{O}_3 + 6\text{H}^+ + 6\text{F}^-$$

Thus, water is a reductant and itself is oxidized to oxygen or ozone. Fluorine is an oxidant and itself is reduced to fluoride ion.

#423543
Topic: Oxidation Number

Passage
Justify that the following reactions are redox reactions:

$$4\text{NH}_3 (g) + 5\text{O}_2 (g) \rightarrow 4\text{NO}(g) + 6\text{H}_2 (s)$$

Solution
The oxidation number of nitrogen increases from -3 to +2. The oxidation number of oxygen decreases from 0 to -2. Ammonia is oxidized and oxygen is reduced. Thus, it is a redox reaction.

---

#423584

**Topic:** Balance redox reactions

Consider the reactions:

Why it is more appropriate to write these reactions as: 

\[ 6 \text{CO}_2(g) + 12\text{H}_2\text{O}(l) \rightarrow \text{C}_6\text{H}_12\text{O}_6(aq) + 6\text{O}_2(g) + 6\text{H}_2\text{O}(l) \]

Also suggest a technique to investigate the path of the above (a) and (b) redox reactions.

\[ 6 \text{CO}_2(g) + 6\text{H}_2\text{O}(l) \rightarrow \text{C}_6\text{H}_12\text{O}_6(aq) + 6\text{O}_2(g) \]

**Solution**

The given reaction is

\[ 6 \text{CO}_2(g) + 6\text{H}_2\text{O}(l) \rightarrow \text{C}_6\text{H}_12\text{O}_6(aq) + 6\text{O}_2(g) \]

It is more appropriate to write this reaction as

\[ 6 \text{CO}_2(g) + 12\text{H}_2\text{O}(l) \rightarrow \text{C}_6\text{H}_12\text{O}_6(aq) + 6\text{O}_2(g) + 6\text{H}_2\text{O}(l) \]

This is because water is produced during photosynthesis and water must be shown on product side. To investigate the path, \( \text{H}_2\text{O}^{18} \) is used instead of \( \text{H}_2\text{O}^{16} \). Thus, instead of using normal O atom in water molecule, we use radioactively labelled O atom in water molecule. By determining the intermediates and products containing labelled oxygen, we can investigate above paths.

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#423618

**Topic:** Balance redox reactions

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 the balanced redox equation for the reaction.

**Solution**

Toluene is oxidized to benzoic acid by using alcoholic potassium permanganate.

In neutral medium hydroxide ions are produced. This reduces cost of adding acid or base.

Alcohol and \( \text{KMnO}_4 \) are both polar and homogeneous to each other. Toluene and alcohol are homogeneous as they are organic. Thus, in alcohol the rate of reaction between toluene and \( \text{KMnO}_4 \) is higher. The balanced redox reaction is as shown.

\[
\begin{align*}
\text{CH}_3\text{C}_6\text{H}_4\text{O}^\cdot & + 2\text{MnO}_4(aq) \rightarrow \\
\text{CO}_2^- & + 2\text{MnO}_2^{2+} + \text{H}_2\text{O}_2(aq) + \text{OH}^- (aq)
\end{align*}
\]

---

#423621

**Topic:** Types of redox reactions

How do you count for the following observations?

When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colourless pungent smell gas \( \text{HCl} \), but if the mixture contains bromide then we get red vapour of bromine. Why?

**Solution**

When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colourless pungent smelling gas \( \text{HCl} \), but if the mixture contains bromide then we get red vapour of bromine. This is because \( \text{HBr} \) can be formed only if dil \( \text{H}_2\text{SO}_4 \) is used. Note: concentrated sulphuric acid converts inorganic chloride to \( \text{HCl} \). But concentrated sulphuric acid converts inorganic bromide to bromine. Dilute sulphuric acid will convert inorganic bromide to \( \text{HBr} \).

---

#423622

**Topic:** Types of redox reactions

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

**Solution**

...
When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colourless pungent smelling gas HCl, but if the mixture contains bromide then we get red vapour of bromine. This is because HBr can be formed only if dil H2SO4 is used.

Note: concentrated sulphuric acid converts inorganic chloride to HCl.
But concentrated sulphuric acid converts inorganic bromide to bromine.
Dilute sulphuric acid will convert inorganic bromide to HBr.

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#423632
Topic: Oxidation Number

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

\[ 2AgBr(s) + C_6H_4O_2(aq) \rightarrow 2Ag(s) + 2HBr(aq) + C_6H_4O_2(aq) \]

Solution
\[ Ag^+ \] is reduced and acts as oxidizing agent.
\[ C_6H_4O_2 \] is oxidized and acts as reducing agent.

#423636
Topic: Oxidation Number

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

\[ HCHO(l) + 2[Ag(NH_3)_2]^+(aq) + 3OH^-(aq) \rightarrow 2Ag(s) + HCOO^-(aq) + 4NH_3(aq) + 2H_2O(l) \]

Solution
\[ HCHO \] is oxidized and \[ [Ag(NH_3)_2]^+ \] is reduced.
\[ [Ag(NH_3)_2]^+ \] is the oxidizing agent and \[ HCHO \] is reducing agent.

#423659
Topic: Oxidation Number

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

\[ HCHO(1) + 2Cu^{2+}(aq) + 5OH^-(aq) \rightarrow Cu_2O(s) + HCOO^-(aq) + 3H_2O(l) \]

Solution
\[ HCHO \] is oxidized and \[ Cu^{2+} \] is reduced.
\[ Cu^{2+} \] is the oxidizing agent and \[ HCHO \] is the reducing agent.

#423661
Topic: Oxidation Number

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

\[ Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l) \]

Solution
Pb is oxidized ad PbO_2 is reduced.
\[ PbO_2 \] is oxidizing agent and Pb is reducing agent.
Balance the following redox reactions by ion electron method.

\[ MnO_4^- (aq) + I^- (aq) \rightarrow MnO_2(s) + I_2(s) \] (in basic medium)
The unbalanced chemical equation is

\[ \text{MnO}_2 (aq) + I^- (aq) \rightarrow \text{MnO}_2 (s) + I_2 (s) \]

The oxidation half reaction is

\[ I^- (aq) \rightarrow I_2 (s) \]

The reduction half reaction is

\[ \text{MnO}_2 (aq) \rightarrow \text{MnO}_2 (aq) \]

Balance I atoms and charges in the oxidation half reaction.

\[ 2I^- (aq) \rightarrow I_2 (s) + 2e^- \]

In the reduction half reaction, the oxidation number of Mn changes from +7 to +2. Hence, add 3 electrons to reactant side of the reaction.

\[ \text{MnO}_2 (aq) + 3e^- \rightarrow \text{MnO}_2 (aq) \]

Balance charge in the reduction half reaction by adding 4 hydroxide ions to product side.

\[ \text{MnO}_2 (aq) + 3e^- + 2H_2O \rightarrow \text{MnO}_2 (aq) + 4OH^- \]

To balance O atoms, add 2 water molecules to reactant side.

\[ \text{MnO}_2 (aq) + 3e^- + 2H_2O \rightarrow \text{MnO}_2 (aq) + 4OH^- \]

To equalize the number of electrons, multiply the oxidation half reaction by 3 and multiply the reduction half reaction by 2.

\[ 6I^- (aq) \rightarrow 3I_2 (s) + 6e^- \]
\[ 2\text{MnO}_2 (aq) = 6e^- + 4H_2O \rightarrow 2\text{MnO}_2 (aq) + 8OH^- \]

Add two half cell reactions to obtain the balanced equation.

\[ 2\text{MnO}_2 (aq) + 6I^- (aq) + 4H_2O (l) \rightarrow 2\text{MnO}_2 (s) + 3I_2 (s) + 8OH^- \]

---

**#423675**

**Topic:** Balance redox reactions

**Passage**

Balance the following redox reactions by ion electron method.

\[ \text{MnO}_2 (aq) + \text{SO}_2 (g) \rightarrow \text{Mn}^{2+} (aq) + \text{HSO}_4^- (aq) \text{[in acidic solution]} \]

**Solution**

The unbalanced chemical equation is:

\[ \text{MnO}_2 (aq) + \text{SO}_2 (g) \rightarrow \text{Mn}^{2+} (aq) + \text{HSO}_4^- (aq) \]

The oxidation half reaction is \( \text{SO}_2 (g) \rightarrow 2\text{H}_2\text{O} (l) \rightarrow \text{HSO}_4^- (aq) + 3\text{H}^+ (aq) + 2e^- (aq) \)

The reduction half reaction is \( \text{MnO}_2 (aq) \rightarrow \text{Mn}^{2+} (aq) \rightarrow \text{Mn}^{2+} (aq) \)

In the reduction half reaction, the oxidation number of Mn changes from +7 to +2. Hence, 5 electrons are added to LHS of the reaction.

\[ \text{MnO}_2 (aq) + 5e^- \rightarrow \text{Mn}^{2+} (aq) \]

Charge is balanced in the reduction half reaction by adding 8 hydrogen ions to LHS.

\[ \text{MnO}_2 (aq) + 5e^- + 8\text{H}^+ (aq) \rightarrow \text{Mn}^{2+} (aq) + 4\text{H}_2\text{O} (l) \]

To balance O atoms, 4 water molecules are added on RHS.

\[ \text{MnO}_2 (aq) + 5e^- + 8\text{H}^+ (aq) \rightarrow \text{Mn}^{2+} (aq) + 4\text{H}_2\text{O} (l) \]

To equalize the number of electrons, the oxidation half reaction is multiplied by 5 and the reduction half reaction is multiplied by 2.

\[ 5\text{SO}_2 (g) + 10\text{H}_2\text{O} (l) + 5\text{H}_2\text{O} (aq) + 15\text{H}^+ (aq) + 10e^- (aq) \]
\[ 2\text{MnO}_2 (aq) = 10e^- + 16\text{H}^+ (aq) \rightarrow 2\text{Mn}^{2+} (aq) \]

Two half cell reactions are added to obtain the balanced equation.

\[ 2\text{MnO}_2 (aq) + 5\text{SO}_2 (g) + 2\text{H}_2\text{O} (l) \rightarrow 2\text{Mn}^{3+} (aq) + 5\text{HSO}_4^- (aq) \]
#423677

**Topic:** Balance redox reactions

**Passage**

Balance the following redox reactions by ion electron method.

\[ H_2O_2 (aq) + Fe^{3+} (aq) \rightarrow Fe^{2+} (aq) + H_2O(l) \] (in acidic solution)

**Solution**

The oxidation half reaction is \( Fe^{3+} (aq) \rightarrow Fe^{2+} (aq) + e^- \).

The reduction half reaction is \( H_2O_2 (aq) = 2H^+ (aq) + 2e^- \rightarrow 2H_2O(l) \)

In above half reactions, all the atoms are balanced.

The oxidation half reaction is multiplied by 2 so as to balance the number of electrons in oxidation half reaction with reduction half reaction.

\[ 2Fe^{3+} (aq) \rightarrow 2Fe^{2+} (aq) + 2e^- \]

The oxidation half reaction is then added to the reduction half reaction to obtain balanced chemical equation.

\[ H_2O_2 (aq) + 2Fe^{3+} (aq) + 2H^+ \rightarrow 2Fe^{2+} (aq) + 2H_2O(l) \]

---

#423679

**Topic:** Balance redox reactions

**Passage**

Balance the following redox reactions by ion electron method.

\[ Cr_2O_7^{2-} + SO_2 (g) \rightarrow Cr^{3+} (aq) + SO_4^{2-} (aq) \] (in acidic solution)

**Solution**

The oxidation half reaction is \( SO_2 (g) = 2H_2O(l) \rightarrow SO_4^{2-} + 4H^+ (aq) + 2e^- \).

The reduction half reaction is \( Cr_2O_7^{2-} (aq) + 14H^+ (aq) + 6e^- \rightarrow 2Cr^{3+} (aq) = 7H_2O(l) \)

The oxidation half reaction is multiplied by 3 and added to the reduction half reaction to obtain the balanced redox reaction.

\[ 3SO_2 (g) + 2H^+ (aq) \rightarrow 2Cr^{3+} (aq) + 3SO_4^{2-} (aq) + H_2O(l) \]

---

#423686

**Topic:** Balance redox reactions

**Passage**

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

\[ P_4 (s) + OH^- (aq) \rightarrow PH_4 (g) + HPO_4^{2-} (aq) \]

**Solution**
Oxidation number method:
The oxidation number of P decreases from 0 to -3 and increases from 0 to +2. Hence, \( P_4 \) is oxidizing as well as reducing agent.
During reduction, the total decrease in the oxidation number for 4 P atoms is 12.
During oxidation, total increase in the oxidation number for 4 P atoms is 4.
The increase in the oxidation number is balanced with decrease in the oxidation number by multiplying \( H_2PO_4^- \) with 3.
\[
P_4(s) + OH^- (aq) \rightarrow PH_3(g) + 3H_2PO_4^-(aq)
\]
To balance O atoms, multiply \( OH^- \) ions by 6.
\[
P_4(s) + 6OH^- (aq) \rightarrow PH_3(g) + 3H_2PO_4^-(aq)
\]
To balance H atoms, 3 water molecules are added to L.H.S and 3 hydroxide ions on R.H.S.
\[
P_4(s) + 6OH^- (aq) + 3H_2O(l) \rightarrow PH_3(g) + 3H_2PO_4^- (aq) + 3OH^- (aq)
\]
Subtract 3 hydroxide ions from both sides.
\[
P_4(s) + 3OH^- (aq) + 3H_2O(l) \rightarrow PH_3(g) + 3H_2PO_4^- (aq)
\]

Ion electron method:
The oxidation half reaction is \( P_4(s) \rightarrow H_2PO_4^- (aq) \).
The P atom is balanced.
\[
P_4(s) \rightarrow 4H_2PO_4^- (aq)
\]
The oxidation number is balanced by adding 4 electrons on RHS.
\[
P_4(s) \rightarrow 4H_2PO_4^- (aq) + 4e^-
\]
The charge is balanced by adding 8 hydroxide ions on LHS.
\[
P_4(s) + 8OH^- (aq) \rightarrow 4H_2PO_4^- (aq)
\]
The O and H atoms are balanced.
The reduction half reaction is \( P_4(s) \rightarrow PH_3(g) \).
The oxidation number is balanced by adding 12 electrons on LHS.
\[
P_4(s) + 12e^- \rightarrow PH_3(g)
\]
The charge is balanced by adding 12 hydroxide ions on RHS.
\[
P_4(s) + 12e^- \rightarrow PH_3(g) + 12OH^-
\]
The oxidation half reaction is multiplied by 3 and the reduction half reaction is multiplied by 2.
The half reactions are then added to obtain balanced chemical equation.
\[
P_4(s) + 3OH^- (aq) + 3H_2O(l) \rightarrow PH_3(g) + 3H_2PO_4^- (aq)
\]
The oxidation number of $N$ increases from $-2$ to $-2$. The oxidation number of $Cl$ decreases from $-5$ to $-1$. Hence, hydrazine is the reducing agent and chlorate ion is the oxidizing agent.

Ion-electron method:
The oxidation half reaction is

$$N_2H_4(l) \rightarrow NO(g)$$

Balance the N atoms.

$$N_2H_4(l) \rightarrow 2NO(g)$$

To balance oxidation number, add 8 electrons.

$$N_2H_4(l) \rightarrow 2NO(g) + 8e^-$$

Add 8 hydroxide ions to balance the charge.

$$N_2H_4(l) + 8OH^- (aq) \rightarrow 2NO(g) + 8e^-$$

The reduction half reaction is

$$ClO_3^- (aq) \rightarrow Cl^-(aq)$$

Add 6 electrons to balance the oxidation number.

$$ClO_3^- (aq) + 6e^- \rightarrow Cl^- (aq)$$

Add 6 hydroxide ions to balance the charge.

$$ClO_3^- (aq) + 6e^- \rightarrow Cl^- (aq) + 6OH^- (aq)$$

Multiply the oxidation half reaction by 3 and multiply the reduction half reaction by 2. Add two half reactions obtain the balanced chemical equation.

$$3N_2H_4(aq) + 4ClO_3^- (aq) \rightarrow 6NO(s) + 4Cl^-(aq) + 6H_2O(aq)$$

Oxidation number method:
Total decrease in oxidation number of $N$ is 8. Total increase in the oxidation number of $Cl$ is 6.

Multiply $N_2H_4$ with 3 and multiply $ClO_3^-$ with 4.

$$3N_2H_4(l) + 4ClO_3^- (aq) \rightarrow NO(g) + Cl^- (aq)$$

Balance N and Cl atoms.

$$3N_2H_4(l) + 4ClO_3^- (aq) \rightarrow 6NO(g) + 4Cl^- (aq)$$

Balance the O atoms by adding 6 water molecules.

$$3N_2H_4(l) + 4ClO_3^- (aq) \rightarrow 6NO(g) + 4Cl^- (aq) + 6H_2O(l)$$

This is the balanced chemical equation.

#423688

**Topic:** Balance redox reactions

**Passage**
Balance the following equations in basic medium by ion-electron method and oxidation number methods and identify the oxidising agent and the reducing agent.
$\text{Cl}_2\text{O}_7^-(g) + \text{H}_2\text{O}_2(aq) \rightarrow \text{ClO}_2^-(aq) + \text{O}_2(g) + \text{H}^+$

### Solution

The oxidation number of chlorine decreases from +7 to +3 and the oxidation number of $\text{O}$ increases from -1 to zero. Thus, $\text{Cl}_2\text{O}_7^-$ is oxidizing agent and $\text{H}_2\text{O}_2$ is the reducing agent.

**Ion electron method:**

The oxidation half equation is

$$\text{H}_2\text{O}_2(aq) \rightarrow \text{O}_2(g) + 2e^-$$

To balance oxidation number, 2 electrons are added.

$$\text{H}_2\text{O}_2(aq) \rightarrow \text{O}_2(g) + 2e^-$$

2 hydroxide ions are added to balance the charge.

$$\text{H}_2\text{O}_2(aq) + 2\text{OH}^- \rightarrow \text{O}_2(g) + 2e^-$$

2 water molecules are added to balance the O atoms.

$$\text{H}_2\text{O}_2(aq) + 2\text{OH}^- \rightarrow \text{O}_2(g) + 2\text{H}_2\text{O}(aq) + 2e^-$$

The reduction half reaction is $\text{Cl}_2\text{O}_7^- \rightarrow \text{ClO}_2^-(aq)$

The Cl atoms are balanced

$\text{Cl}_2\text{O}_7^- \rightarrow \text{ClO}_2^-(aq)$

8 electrons are added to balance the oxidation number.

$\text{Cl}_2\text{O}_7^- + 8e^- \rightarrow \text{ClO}_2^-(aq)$

6 hydroxide ions are added to balance the charge.

$\text{Cl}_2\text{O}_7^- + 8e^- \rightarrow \text{ClO}_2^-(aq) + 6\text{OH}^-(aq)$

The oxidation half equation is multiplied with 4 and added to reduction half equation.

$4\text{Cl}_2\text{O}_7^-(g) + 4\text{H}_2\text{O}_2(aq) + 2\text{OH}^- \rightarrow 4\text{ClO}_2^-(aq) + 4\text{O}_2(g) + 5\text{H}_2\text{O}(l)$

**Oxidation number method:**

Total decrease in oxidation number of $\text{Cl}_2\text{O}_7^-$ is 8.

Total increase in oxidation number of $\text{H}_2\text{O}_2$ is 2.

$\text{H}_2\text{O}_2$ and $\text{O}_2$ are multiplied with 4

$4\text{Cl}_2\text{O}_7^-(g) + 4\text{H}_2\text{O}_2(aq) \rightarrow 4\text{ClO}_2^-(aq) + 4\text{O}_2(g)$

Chlorine atoms are balanced

$4\text{Cl}_2\text{O}_7^-(g) + 4\text{H}_2\text{O}_2(aq) \rightarrow 2\text{ClO}_2^-(aq) + 4\text{O}_2(g)$

$\text{O}$ atoms are balanced by adding 3 water molecules.

$4\text{Cl}_2\text{O}_7^-(g) + 4\text{H}_2\text{O}_2(aq) \rightarrow 2\text{ClO}_2^-(aq) + 4\text{O}_2(g) + 3\text{H}_2\text{O}(l)$

$\text{H}$ atoms are balanced by adding 2 hydroxide ions and 2 water molecules.

$4\text{Cl}_2\text{O}_7^-(g) + 4\text{H}_2\text{O}_2(aq) + 2\text{OH}^- \rightarrow 2\text{ClO}_2^-(aq) + 4\text{O}_2(g) + 5\text{H}_2\text{O}(l)$

### #423689

**Type:** Types of redox reactions

What sorts of informations can you draw from the following reaction?

$(\text{CN})_2(g) + 2\text{OH}^-(aq) \rightarrow \text{CN}^-(aq) + \text{CNO}^-(aq) + \text{H}_2\text{O}(l)$

**Solution**
The reaction is a disproportionation reaction.
It occurs in basic medium.
The oxidation number of N in \((\text{CN})_2\), \(\text{CN}^-\) and \(\text{CNO}^-\) is \(-3\), \(-2\) and \(-5\) respectively.
Cyanogen \((\text{CN})_2\) is simultaneously reduced to \(\text{CN}^-\) ion and oxidised to cyanate ion \(\text{CNO}^-\) ion.

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**#423690**

**Topic:** Balance redox reactions

The \(\text{Mn}^{3+}\) ion is unstable in solution and undergoes disproportionation to give \(\text{Mn}^{2+}\), \(\text{MnO}_2\) and \(\text{H}^+\) ion. Write a balanced ionic equation for the reaction.

**Solution**

The unbalanced chemical reaction is:
\[
\text{Mn}^{3+}\ (aq) \rightarrow \text{Mn}^{2+}\ (aq) + \text{MnO}_2\ (s) + \text{H}^+\ (aq)
\]

The oxidation half reaction is:
\[
\text{Mn}^{3+}\ (aq) \rightarrow \text{MnO}_2\ (s)
\]

To balance oxidation number, one electron is added on R.H.S.
\[
\text{Mn}^{3+}\ (aq) \rightarrow \text{MnO}_2\ (s) + e^-
\]

4 protons are added to balance the charge.
\[
\text{Mn}^{3+}\ (aq) \rightarrow \text{MnO}_2\ (s) + 4\text{H}^+\ (aq) + e^-
\]

2 water molecules are added to balance O atoms.

The reduction half reaction is \(\text{Mn}^{3+}\ (aq) \rightarrow \text{Mn}^{2+}\ (aq)\).

An electron is added to balance oxidation number.
\[
\text{Mn}^{3+}\ (aq) + e^- \rightarrow \text{Mn}^{2+}\ (aq)
\]

Two half cell reactions are added to obtain balanced chemical equation.
\[
2\text{Mn}^{3+}\ (aq) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{Mn}^{2+}\ (aq) + \text{MnO}_2\ (s) + 4\text{H}^+\ (aq)
\]

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**#423696**

**Topic:** Balance redox reactions

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 the redox change taking place in water.

**Solution**

The balanced chemical reaction for the redox reaction between chlorine and sulphur dioxide is:
\[
\text{Cl}_2 + \text{SO}_2 + 2\text{H}_2\text{O} \rightarrow 2\text{Cl}^- + \text{SO}_4^{2-} + 4\text{H}^+
\]

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**#423698**

**Topic:** Types of redox reactions

**Passage**

Refer to the periodic table given in your book and now answer the following questions:

Select three possible non metals that can show disproportionation reaction.

**Solution**

Phosphorus, chlorine and sulphur are the non metals which can show disproportionation reaction. The reactions shown by them are:
\[
\text{P}_4 + 3\text{OH}^- + \text{3H}_2\text{O} \rightarrow \text{PH}_3 + 3\text{H}_3\text{PO}_2
\]
\[
\text{Cl}_2 + 2\text{OH}^- \rightarrow \text{Cl}^- + \text{ClO}^- + \text{H}_2\text{O}
\]
\[
\text{S}_8 + 12\text{OH}^- \rightarrow 4\text{S}^{2-} + 2\text{SO}_4^{2-} + 6\text{H}_2\text{O}
\]

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**#423699**

**Topic:** Types of redox reactions
Passage
Refer to the periodic table given in your book and now answer the following questions:

Select three metals that can show disproportionation reaction.

Solution
Copper, gallium and indium are the metals that show disproportionation reaction. The reactions are shown below.

\[ 2Cu^+ \rightarrow Cu^2+ + Cu \]
\[ 3Ga^+ \rightarrow Ga^3+ + 2Ga \]
\[ 3In^+ \rightarrow In^{3+} + 2In \]

#423724
Topic: Electrode potential

Passage
Depict the galvanic cell in which the reaction \( \text{Zn}(s) + 2 Ag^+(aq) \rightarrow \text{Zn}^{2+}(aq) + 2 Ag(s) \) takes place. Further show:

the carriers of the current in the cell.

Solution
For the given redox reaction, the galvanic cell is

\[ \text{Zn} | \text{Zn}^{2+}(aq) || \text{Ag}^+(aq) | \text{Ag} \]

The current is carried by the ions in the cell.